Edrolo

VCE CHEMISTRY

Units 1 & 2

Thomas Koster, Giao Nguyen, Chris Sparks,
Thomas Sutherland, Katherine Roan

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PUBLISHED IN AUSTRALIA BY Edrolo
144 George Street Fitzroy VIC 3065, Australia

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National Library of Australia Cataloguing-in-Publication data

TITLE: Edrolo VCE Chemistry Units 1 & 2
CREATOR: Edrolo et al.
ISBN: 978-0-6489433-2-7
TARGET AUDIENCE: For secondary school age.
SUBJECTS: Chemistry--Study and teaching (Secondary)--Victoria
Chemistry--Victoria--Textbooks.
Chemistry--Theory, exercises, etc.
OTHER CREATORS/CONTRIBUTORS: Alana Komesaroff, Rebecca Komesaroff, Domenica Trifilo, Ateeshnum Chandra, Marnie Bates, Miriam Beasy

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TYPE DESIGN: Deborah Johnson, Nathan McGinness, Pamela Sicari, Christopher Dell
TYPESET BY Christopher Dell, Dean Dragonetti, Emma Wright
COVER DESIGN BY Deborah Johnson

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CONTENTS

USING THIS RESOURCE TO TEACH AND LEARN VI
FEATURES OF THIS BOOK VIII

01 **Scientific investigation** 1

1A Scientific research 2
1B Conducting an experiment 9
1C Interpreting data 18
1D Experimental factors affecting data 26
1E Writing scientific material 36
Chapter 1 review 45

UNIT 1: HOW CAN THE DIVERSITY OF MATERIALS BE EXPLAINED? **49**

AREA OF STUDY 1: HOW CAN KNOWLEDGE OF ELEMENTS EXPLAIN THE PROPERTIES OF MATTER? 50

02 **Atoms, particles and molecules** 51

2A The atom 52
2B Defining elements 66
2C The periodic table - part 1 72
2D The periodic table - part 2 81
2E Relative masses 88
2F The mole - part 1 98
2G The mole - part 2 104
2H Chemical formulas and equations 113
Chapter 2 review 126

03 **Metals** 131

3A Metals 132
3B Variations of metallic substances 142
3C Reactivity of metals 150
Chapter 3 review 158

04 **Ionic compounds** 163

4A Ionic compounds 164

4B Properties of ionic compounds 171

4C Writing formulas for ionic compounds 178

Chapter 4 review 188

AREA OF STUDY 2: HOW CAN THE VERSATILITY OF NON-METALS BE EXPLAINED? 194

05 **Molecular theory** 195

5A Covalent bonding 196

5B Properties of covalent compounds 206

5C Intermolecular bonding 216

Chapter 5 review 227

06 **Organic chemistry** 233

6A Carbon materials 234

6B Introduction to hydrocarbons 246

6C Naming simple hydrocarbons 259

6D Naming complex hydrocarbons 273

6E Functional groups - alcohols and carboxylic acids 281

6F Functional groups - esters 291

6G Polymers 299

6H Modifying polymers 307

Chapter 6 review 316

UNIT 2: WHAT MAKES WATER SUCH A UNIQUE CHEMICAL? 321

AREA OF STUDY 1: HOW DO SUBSTANCES INTERACT WITH WATER? 322

07 **Water** 323

7A Properties of water 324

7B Thermal properties of water 335

7C Dissolution 344

7D Concentration 352

7E Solubility 364

7F Precipitation reactions 376

Chapter 7 review 385

08 **Acids and bases** 391

8A An introduction to acids and bases 392

8B Measuring acidity 407

8C Reactions of acids 420

8D Acid-base chemistry in the body 430

Chapter 8 review 439

09 **Redox reactions** 447

9A An introduction to redox reactions 448

9B Redox equations 458

9C Metal displacement reactions 472

9D Redox chemistry in the body 481

Chapter 9 review 491

AREA OF STUDY 2: HOW ARE SUBSTANCES IN WATER MEASURED AND ANALYSED? 498

10 **Analysis of solutions** 499

10A Water on earth 500

10B Water contamination 505

10C Analysing water using gravimetric analysis 516

10D Analysing water using light - part 1 528

10E Analysing water using light - part 2 538

10F Volumetric analysis 546

10G Chromatography 563

Chapter 10 review 573

Answers 579

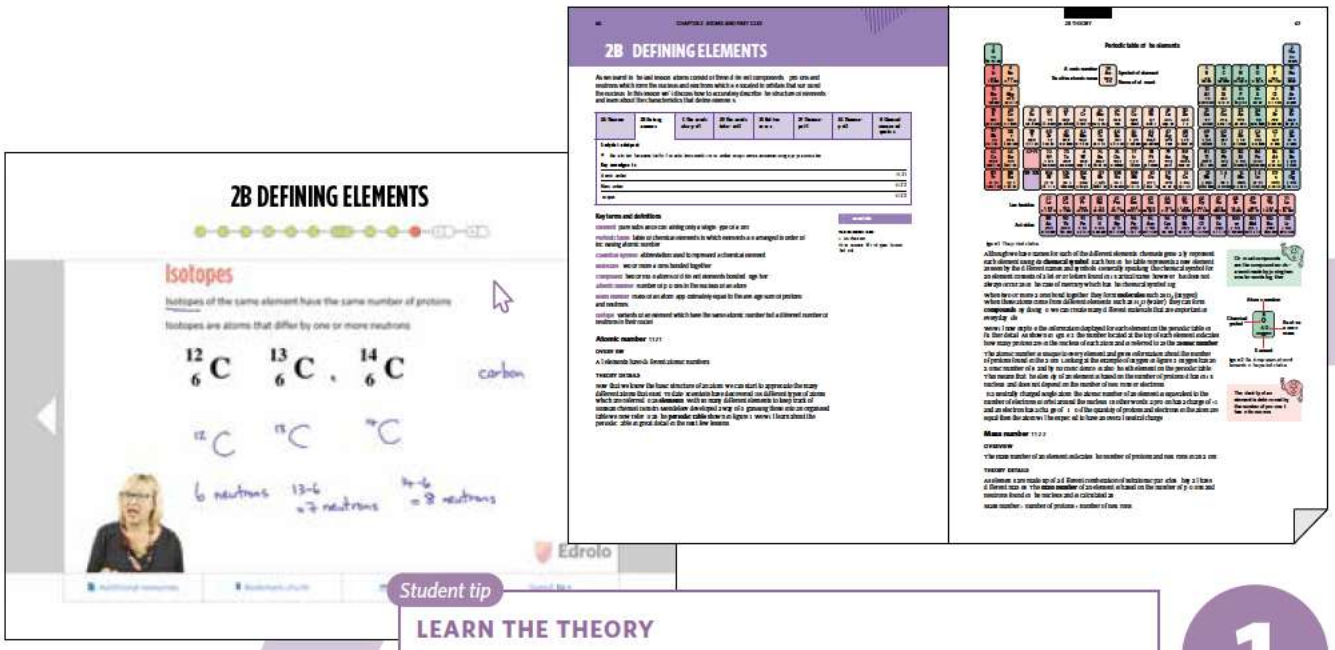
GLOSSARY 682

ACKNOWLEDGEMENTS 690

RESEARCH QUESTIONS FOR UNIT 1 AOS 3 691

VCAA DATA BOOK 693

USING THIS RESOURCE TO TEACH AND LEARN



Student tip

LEARN THE THEORY

Every dot-point in your study design is covered in our video lessons and textbook theory – perfect to use for pre-learning, during class, and as revision.

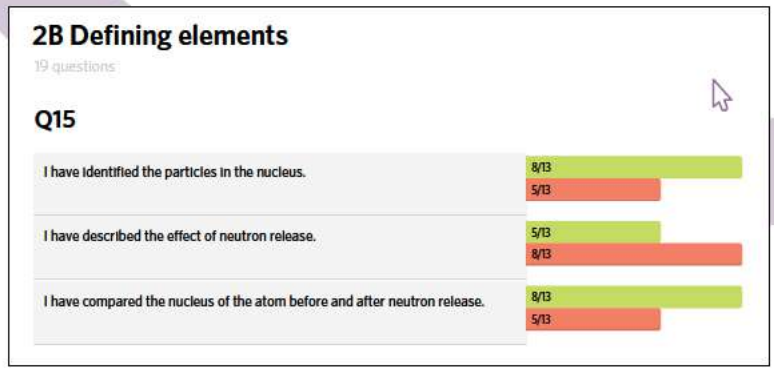
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Teacher tip

EVALUATE STRENGTHS AND AREAS FOR IMPROVEMENT

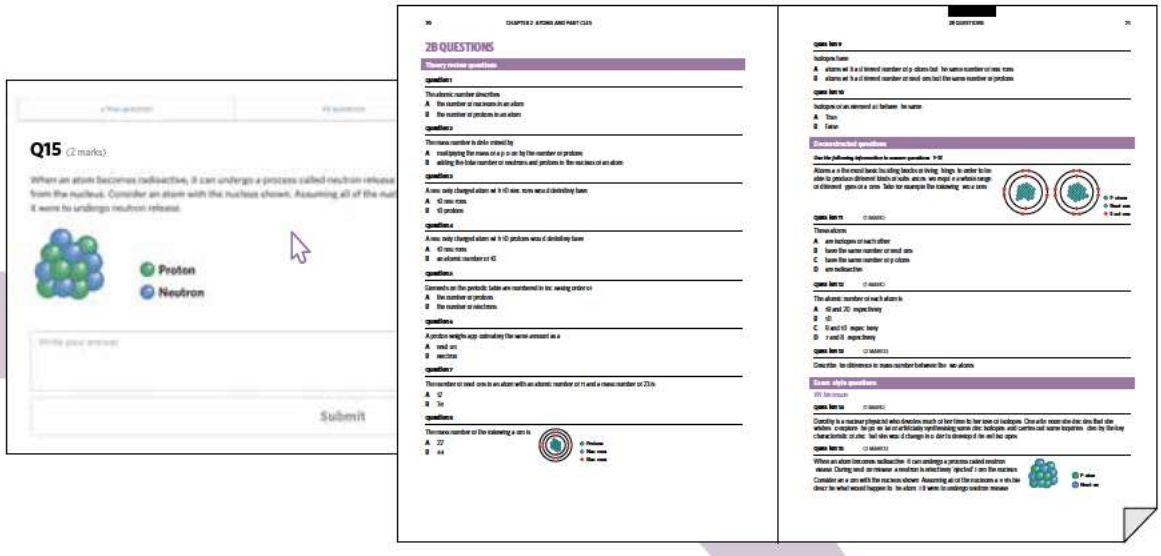
Teachers see class-level data and individual student responses – use this to provide feedback, differentiate student learning, plan future lessons, and inform the revision program of your students.

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Student tip

CHECK FOR UNDERSTANDING

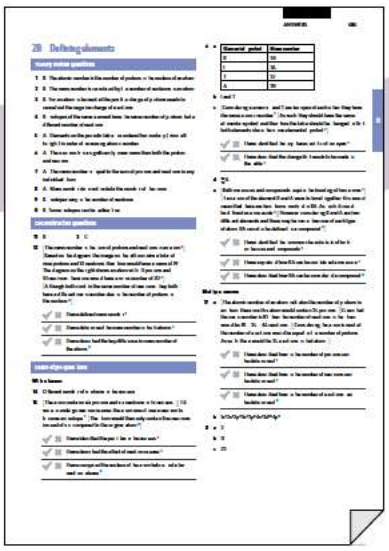
Each lesson has theory review questions, a deconstructed exam-style question, and exam-style questions so you can apply your knowledge in different ways and consolidate your learning. You'll also find tests/exams within each area of study.

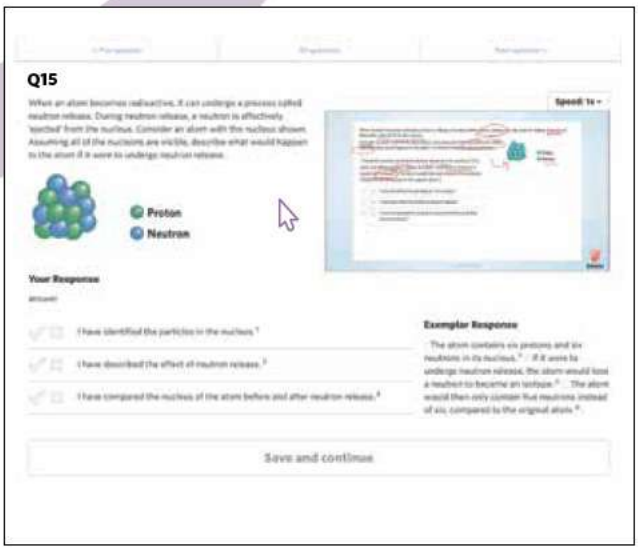
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Student tip

SELF-ASSESS AND GET FEEDBACK

At the back of your textbook you'll find exemplar responses and checklists for every exam-style question. In your Edrolo account, you'll find video solutions as well as the interactive checklists and exemplar responses. Use these answers to target your revision and get the greatest impact from your study time. This enables you to focus on the parts of the theory you struggled with, and ask your teacher for support if you get totally stuck!





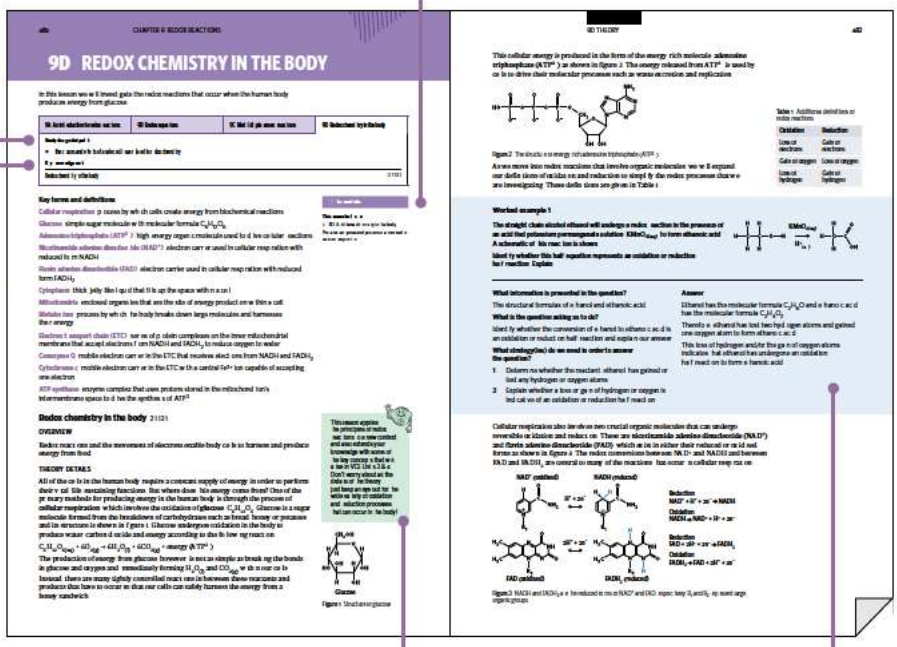
FEATURES OF THIS BOOK

Edrolo's VCE Chemistry Units 1 & 2 textbook has the following features.

Lesson links show how the knowledge in the lesson links with other key concepts in previous lessons.

Study design dot points from the VCAA curriculum provide explicit links between our lessons and the syllabus.

Key knowledge units break down the theory into smaller chunks and can be used to help navigate the corresponding theory lesson videos online.

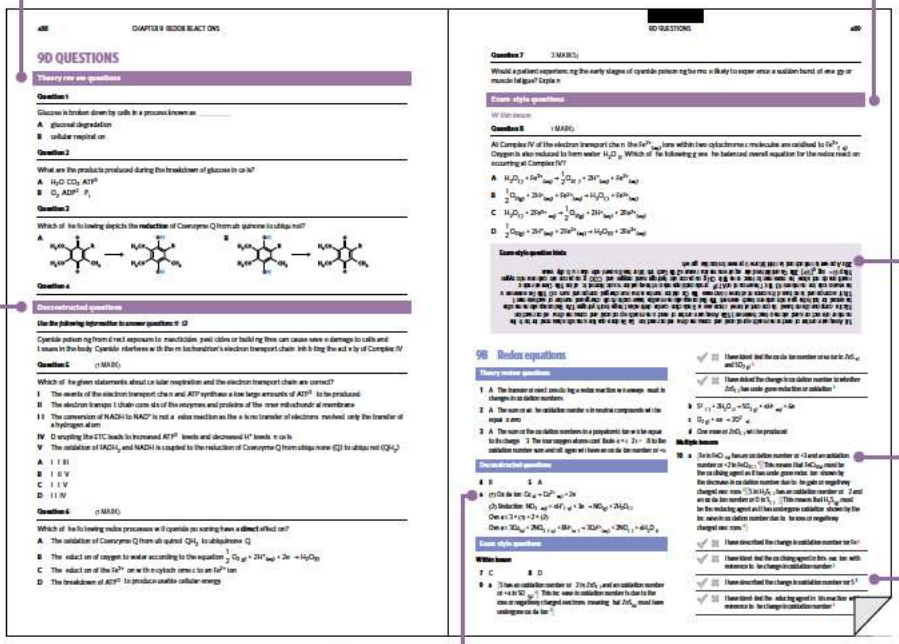


Proton/neutron/electron tips address common misconceptions and/or areas that students often forget that are important in developing their understanding.

Worked examples apply theory to exam-style questions with full working and explanations of each step in the process.

Theory review questions test if students can remember the basic theory and overcome common misconceptions. They are stepping stones between the content and exam-style questions.

Exam-style questions reflect the style of your end-of-year exam in Year 12. These include questions from within the lesson, from previous lessons, and testing key science skills in the context of the theory you just learned.



Exam-style question hints give students a 'starting point' when they attempt exam-style questions when they are unsure.

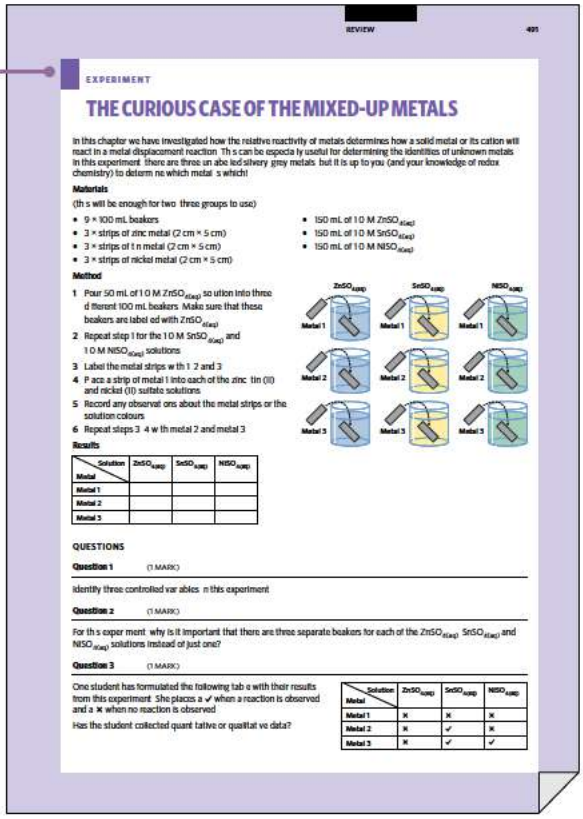
Worded responses and visual responses include an exemplar answer, to show students what a full mark response looks like, as well as a checklist.

Checklists are provided for all worded responses and visual responses. They explain the purpose of each part of a full-mark answer to help students develop their scientific communication skills.

Deconstructed questions provide prompts for students to piece together the steps needed to answer an exam-style question.

Calculation responses show full working to help students form good habits and to provide more detailed guidance for students.

An **experiment** on a topic related to the chapter, accompanied by a set of questions on the experimental investigation relating to that experiment.



REVIEW 491

EXPERIMENT

THE CURIOUS CASE OF THE MIXED-UP METALS

In this chapter we have investigated how the relative reactivity of metals determines how a solid metal or its cation will react in a metal displacement reaction. This can be especially useful for determining the identities of unknown metals. In this experiment, there are three unlabelled silvery grey metals, but it is up to you (and your knowledge of redox chemistry) to determine which metal is which!

Materials

(This will be enough for two–three groups to use)

- 9 × 100 mL beakers
- 3 × strips of zinc metal (2 cm × 5 cm)
- 3 × strips of tin metal (2 cm × 5 cm)
- 3 × strips of nickel metal (2 cm × 5 cm)
- 150 mL of 1.0 M $ZnSO_{4(aq)}$
- 150 mL of 1.0 M $SnSO_{4(aq)}$
- 150 mL of 1.0 M $NiSO_{4(aq)}$

Method

1 Pour 50 mL of 1.0 M $ZnSO_{4(aq)}$ solution into three different 100 mL beakers. Make sure that these beakers are labelled with $ZnSO_{4(aq)}$.
2 Repeat step 1 for the 1.0 M $SnSO_{4(aq)}$ and 1.0 M $NiSO_{4(aq)}$ solutions.
3 Label the metal strips with 1, 2 and 3.
4 Place a strip of metal 1 into each of the zinc, tin (II) and nickel (II) sulfate solutions.
5 Record any observations about the metal strips or the solution colours.
6 Repeat steps 3–4 with metal 2 and metal 3.

Results

Solution / Metal	$ZnSO_{4(aq)}$	$SnSO_{4(aq)}$	$NiSO_{4(aq)}$
Metal 1			
Metal 2			
Metal 3			

QUESTIONS

Question 1 (1 MARK)

Identify three controlled variables in this experiment.

Question 2 (1 MARK)

For this experiment, why is it important that there are three separate beakers for each of the $ZnSO_{4(aq)}$, $SnSO_{4(aq)}$ and $NiSO_{4(aq)}$ solutions instead of just one?

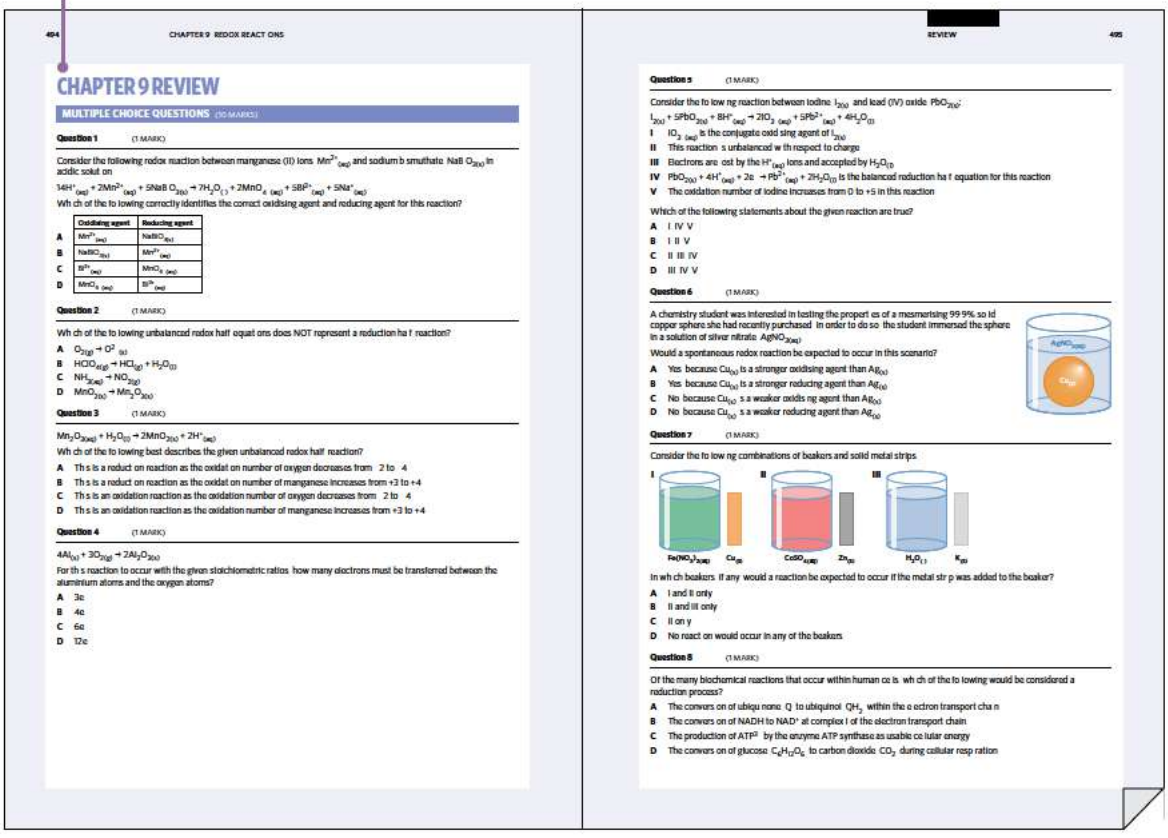
Question 3 (1 MARK)

One student has formulated the following table with their results from this experiment. She places a ✓ when a reaction is observed and a ✗ when no reaction is observed.

Has the student collected quantitative or qualitative data?

Solution / Metal	$ZnSO_{4(aq)}$	$SnSO_{4(aq)}$	$NiSO_{4(aq)}$
Metal 1	✗	✗	✗
Metal 2	✗	✓	✗
Metal 3	✗	✓	✓

The **chapter review** consists of exam-style questions that cover the content from that chapter or area of study. They are structured like the end-of-year exam in Year 12, consisting of a Section A with multiple choice questions and a Section B with short answer questions.



494 CHAPTER 9 REDOX REACTIONS

CHAPTER 9 REVIEW

MULTIPLE CHOICE QUESTIONS (30 MARKS)

Question 1 (1 MARK)

Consider the following redox reaction between manganese (II) ions, $Mn^{2+}_{(aq)}$, and sodium bismuthate, $NaBiO_{3(s)}$, in acidic solution.

$14H^+_{(aq)} + 2Mn^{2+}_{(aq)} + 5NaBiO_{3(s)} \rightarrow 7H_2O_{(l)} + 2MnO_{4\ (aq)}^- + 5Bi^{2+}_{(aq)} + 5Na^+_{(aq)}$

Which of the following correctly identifies the correct oxidising agent and reducing agent for this reaction?

	Oxidising agent	Reducing agent
A	$Mn^{2+}_{(aq)}$	$NaBiO_{3(s)}$
B	$NaBiO_{3(s)}$	$Mn^{2+}_{(aq)}$
C	$Bi^{2+}_{(aq)}$	$MnO_{4\ (aq)}^-$
D	$MnO_{4\ (aq)}^-$	$Bi^{2+}_{(aq)}$

Question 2 (1 MARK)

Which of the following unbalanced redox half-equations does NOT represent a reduction half-reaction?

A $O_{2(g)} \rightarrow O^{2-}_{(s)}$
B $HClO_{(aq)} \rightarrow HCl_{(g)} + H_2O_{(l)}$
C $NH_{3(aq)} \rightarrow NO_{2(g)}$
D $MnO_{2(s)} \rightarrow Mn_2O_{3(s)}$

Question 3 (1 MARK)

$Mn_2O_{3(aq)} + H_2O_{(l)} \rightarrow 2MnO_{2(s)} + 2H^+_{(aq)}$

Which of the following best describes the given unbalanced redox half-reaction?

A This is a reduction reaction as the oxidation number of oxygen decreases from −2 to −4.
B This is a reduction reaction as the oxidation number of manganese increases from +3 to +4.
C This is an oxidation reaction as the oxidation number of oxygen decreases from −2 to −4.
D This is an oxidation reaction as the oxidation number of manganese increases from +3 to +4.

Question 4 (1 MARK)

$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$

For this reaction to occur with the given stoichiometric ratios, how many electrons must be transferred between the aluminium atoms and the oxygen atoms?

A 3e⁻
B 4e⁻
C 6e⁻
D 12e⁻

REVIEW 495

Question 5 (1 MARK)

Consider the following reaction between iodine, $I_{2(s)}$, and lead (IV) oxide, $PbO_{2(s)}$:

$I_{2(s)} + 5PbO_{2(s)} + 8H^+_{(aq)} \rightarrow 2IO_{3\ (aq)}^- + 5Pb^{2+}_{(aq)} + 4H_2O_{(l)}$

I $IO_{3\ (aq)}^-$ is the conjugate oxidising agent of $I_{2(s)}$.
II This reaction is unbalanced with respect to charge.
III Electrons are lost by the $H^+_{(aq)}$ ions and accepted by $H_2O_{(l)}$.
IV $PbO_{2(s)} + 4H^+_{(aq)} + 2e^- \rightarrow Pb^{2+}_{(aq)} + 2H_2O_{(l)}$ is the balanced reduction half-equation for this reaction.
V The oxidation number of iodine increases from 0 to +5 in this reaction.

Which of the following statements about the given reaction are true?

A I, IV, V
B I, II, V
C II, III, IV
D III, IV, V

Question 6 (1 MARK)

A chemistry student was interested in testing the properties of a mesmerising 99.9% solid copper sphere she had recently purchased. In order to do so, the student immersed the sphere in a solution of silver nitrate, $AgNO_{3(aq)}$.

Would a spontaneous redox reaction be expected to occur in this scenario?

A Yes, because $Cu_{(s)}$ is a stronger oxidising agent than $Ag_{(s)}$.
B Yes, because $Cu_{(s)}$ is a stronger reducing agent than $Ag_{(s)}$.
C No, because $Cu_{(s)}$ is a weaker oxidising agent than $Ag_{(s)}$.
D No, because $Cu_{(s)}$ is a weaker reducing agent than $Ag_{(s)}$.

Question 7 (1 MARK)

Consider the following combinations of beakers and solid metal strips.

In which beakers, if any, would a reaction be expected to occur if the metal strip was added to the beaker?

A I and II only
B II and III only
C II only
D No reaction would occur in any of the beakers.

Question 8 (1 MARK)

Of the many biochemical reactions that occur within human cells, which of the following would be considered a reduction process?

A The conversion of ubiquinone, Q, to ubiquinol, QH_2, within the electron transport chain
B The conversion of NADH to NAD^+ at complex I of the electron transport chain
C The production of ATP^{3-} by the enzyme ATP synthase as usable cellular energy
D The conversion of glucose, $C_6H_{12}O_6$, to carbon dioxide, CO_2, during cellular respiration

CHAPTER 1

Scientific investigation

01

1A Scientific research

1B Conducting an experiment

1C Interpreting data

1D Experimental factors affecting data

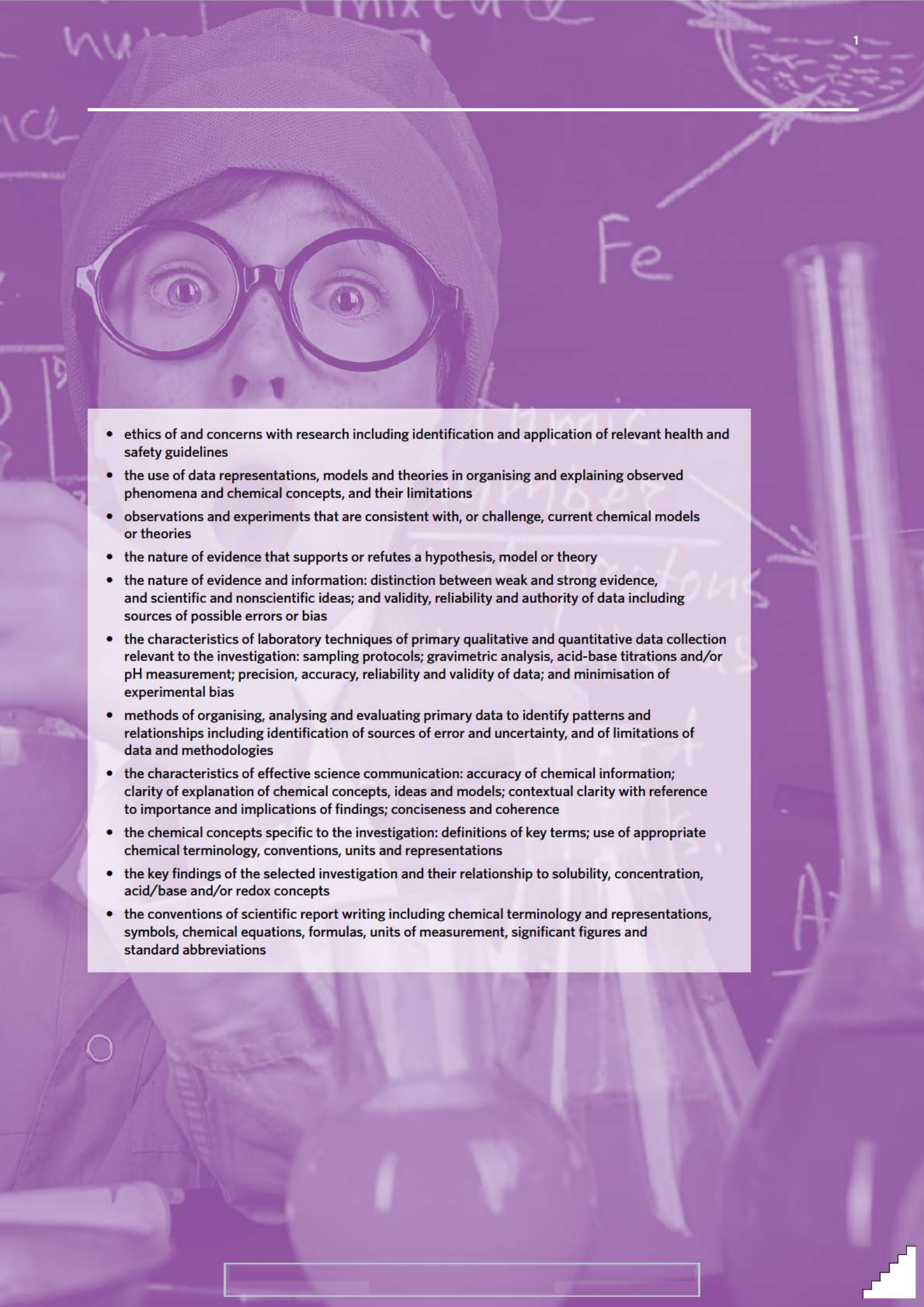
1E Writing scientific material

The development of a set of key science skills is a core component of the study of VCE Chemistry and applies across Units 1 to 4 in all areas of study. This chapter merges the details of AOS3 from both Units 1 and 2 to build the foundation of knowledge required to both analyse and conduct research investigations. The knowledge and skills acquired in this chapter will help you throughout all of your Chemistry studies.

Key knowledge

- the nature of evidence and information: distinction between weak and strong evidence, and scientific and non-scientific ideas; and validity, reliability and authority of data including sources of possible errors or bias
- the chemical concepts specific to the investigation and their significance, including definitions of key terms, and chemical representations
- the characteristics of laboratory techniques of primary qualitative and quantitative data collection relevant to the investigation: sampling protocols; gravimetric analysis, acid-base titrations and/or pH measurement; precision, accuracy, reliability and validity of data; and minimisation of experimental bias
- ethics of and concerns with research including identification and application of relevant health and safety guidelines
- the influence of social, economic, environmental and ethical factors relevant to the selected chemical investigation

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- ethics of and concerns with research including identification and application of relevant health and safety guidelines
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations
- observations and experiments that are consistent with, or challenge, current chemical models or theories
- the nature of evidence that supports or refutes a hypothesis, model or theory
- the nature of evidence and information: distinction between weak and strong evidence, and scientific and nonscientific ideas; and validity, reliability and authority of data including sources of possible errors or bias
- the characteristics of laboratory techniques of primary qualitative and quantitative data collection relevant to the investigation: sampling protocols; gravimetric analysis, acid-base titrations and/or pH measurement; precision, accuracy, reliability and validity of data; and minimisation of experimental bias
- methods of organising, analysing and evaluating primary data to identify patterns and relationships including identification of sources of error and uncertainty, and of limitations of data and methodologies
- the characteristics of effective science communication: accuracy of chemical information; clarity of explanation of chemical concepts, ideas and models; contextual clarity with reference to importance and implications of findings; conciseness and coherence
- the chemical concepts specific to the investigation: definitions of key terms; use of appropriate chemical terminology, conventions, units and representations
- the key findings of the selected investigation and their relationship to solubility, concentration, acid/base and/or redox concepts
- the conventions of scientific report writing including chemical terminology and representations, symbols, chemical equations, formulas, units of measurement, significant figures and standard abbreviations

1A SCIENTIFIC RESEARCH

In this lesson, we will learn about the development of scientific ideas.

1A Scientific research	1B Conducting an experiment	1C Interpreting data	1D Experimental factors affecting data	1E Writing scientific material

Study design dot points

- the nature of evidence and information: distinction between weak and strong evidence, and scientific and non-scientific ideas; and validity, reliability and authority of data including sources of possible errors or bias
- the chemical concepts specific to the investigation and their significance, including definitions of key terms, and chemical representations

Key knowledge units

Scientific vs non-scientific ideas	1.3.4.2
Linking chemical concepts to investigations	2.3.1.1

Key terms and definitions

Scientific method procedure used to investigate scientific ideas

Scientific ideas ideas that are developed by following the scientific method

Non-scientific ideas ideas that are not developed by following the scientific method

Primary source source that provides the original raw data

Secondary source source that has interpreted primary sources

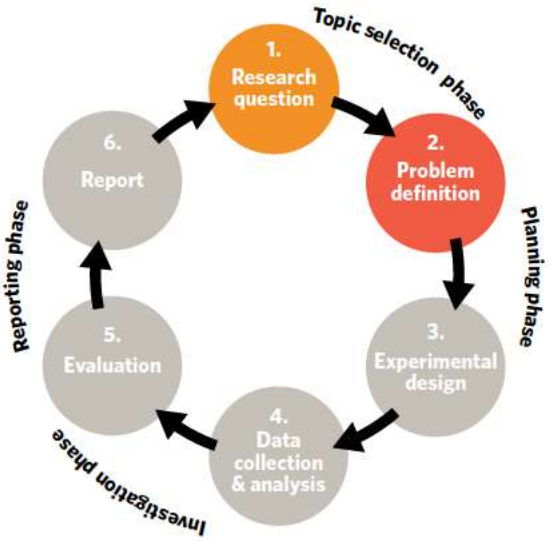
Aim purpose of an experiment

Hypothesis testable statement which predicts the outcome of an experiment

Independent variable variable that is deliberately manipulated by the experimenter

Dependent variable variable that is measured by the experimenter

Controlled variable(s) variable(s) held constant throughout the experiment



Scientific vs non-scientific ideas 1.3.4.2

OVERVIEW

Although we have gained a lot of knowledge over time, only some of that knowledge is considered scientific.

THEORY DETAILS

Throughout history, humans have been able to gain a lot of knowledge about the world and about ourselves. In early history, discoveries and knowledge were passed on to generations through different methods such as storytelling, dancing, and artefacts.

Some of the early knowledge was gained through experiences, intuition and the senses. Over time, we developed a more methodical way of developing and explaining new knowledge; a process referred to as the **scientific method**. This method is outlined in the flowchart found on the previous page.

The scientific method involves conducting experiments in a way that allows us to be confident about the conclusions that we make, and therefore the pieces of knowledge that we gain. It requires us to develop predictions and conduct experiments in a controlled environment to collect information that can later be analysed. From this, we are able to gain a better understanding of the world.

One of the key features of the scientific method is that it needs to be reproducible; that is, others should be able to follow the same steps and achieve the same outcomes. The ideas that are developed using the scientific method are known as **scientific ideas**. The theories that we will study throughout this course have been developed through this process, and can therefore be considered as scientific ideas.

Scientific ideas are presented in a particular format as shown in figure 1. This general structure is followed by all scientists who conduct research and is how scientific journal articles are presented. Developments in scientific knowledge are published in journals and some journals, also known as peer-reviewed journals, are checked by other scientists to ensure that the right processes have been followed and that the conclusions are true. Ultimately, all of this knowledge contributes to our understanding of the world as shown in figure 2.

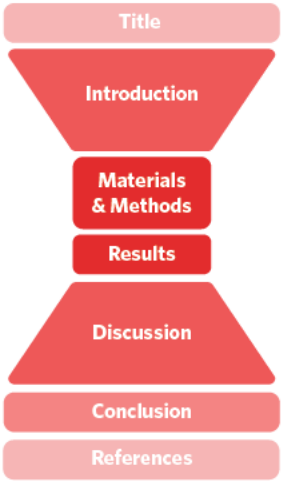


Figure 1 General format used when presenting scientific ideas.

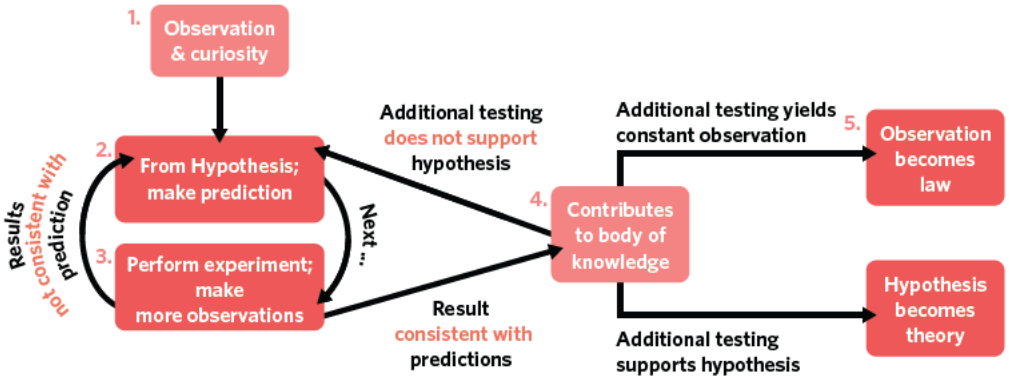


Figure 2 The development of scientific knowledge.

There are some ideas that are hard to measure and therefore have not been developed using the scientific method. We call these ideas **non-scientific ideas**. For example, ideas developed through personal experience or tradition that have not been examined can be considered non-scientific.

Linking chemical concepts to investigations 2.3.1.1

OVERVIEW

Scientific knowledge is built upon previous ideas.

THEORY DETAILS

What we know today about chemistry is a result of all of the knowledge that has been developed previously. The discoveries that we make now build upon this knowledge. Therefore, when choosing a research investigation, we need to acknowledge this by linking it to relevant chemical concepts. In doing so, we can ensure that the basis of our research stems from a solid theoretical foundation.

Brainstorming an investigation

When brainstorming ideas about a specific scientific investigation to pursue, it's useful to choose an area of interest and then conduct background research into what has been already established in the area. Most of the time, this involves finding journal articles of previous research that have investigated questions in the same area/field. By doing this, we are able to gain a better understanding of the current situation, and it also helps to give direction to the kind of research that we can conduct.

For example, say we were interested in investigating water quality. Due to the fact that water is a very important resource, we need to ensure that the water that is available is able to be used for the many different functions in our daily lives. As a result, we need to conduct regular tests to ensure that our water is not contaminated with harmful chemicals and that it has low levels of heavy metals and salts. Figure 3 shows how a research idea can be developed.



Figure 3 Brainstorming of potential research topics.

From this, we can start researching the area to find out more about what is currently understood about the topic as shown in figure 4.

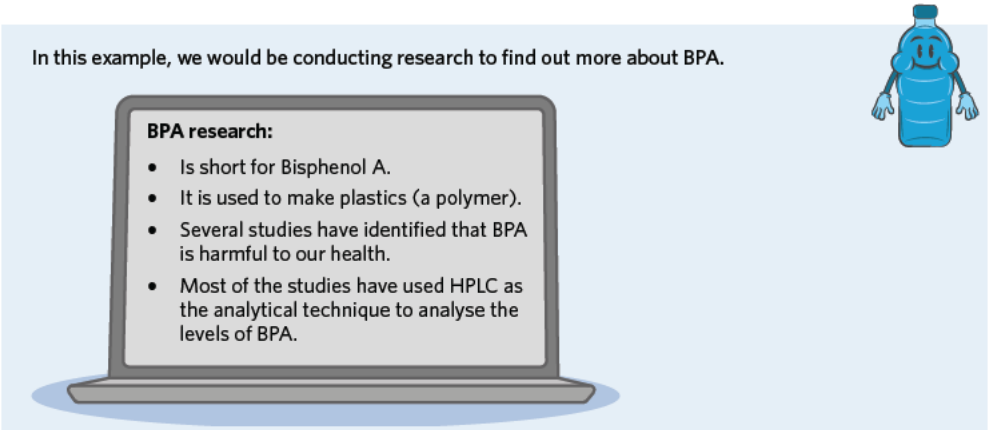


Figure 4 Researching BPA.

The best place to gather information about the chemical theory related to different concepts is through reading journal articles and literature reviews related to the topic of interest. These can be original research articles which contain the experimental data and are referred to as **primary sources**, or other resources such as articles or books that have already reviewed and summarised the data and are referred to as **secondary sources**. These types of articles provide a broad overview of the related theory and include related research that can be further explored depending on the specific area of interest. This step is particularly important as it provides us with:

- relevant terminology and definitions that we will need to use throughout the investigation.
- underlying chemical concepts related to the research.
- possible methods that can be adapted to our own research.

During this process, it is important to make sure that we use resources that have come from reputable sources to ensure that the information is accurate. This mainly involves asking questions such as:

- Who wrote the article?
  - Are they an expert in the field?
- Where was the article published?
  - Was the article published in a peer-reviewed journal?
- Does the author have affiliations with the relevant industry?
  - Have they been funded by a private company in the industry?

Reviewing scientific research

All scientific research is presented in the format as outlined in figure 1. When reviewing the research, there are some key features that we need to consider.

Although we can get a general overview of the theory and purpose of the study from the introduction, the more important sections of scientific research are the results and methods. By interpreting the data presented in the results section, we are able to further understand the theory being tested. The results can support the theory in that it highlights an expected relationship, or it could question the theory itself by showing results that do not match the current theoretical understanding. In situations where the results deviate from the theory, it is important that we are able to evaluate the methodology of the experiment to see whether or not errors may have affected the results. When reviewing the methodology, we need to ask questions like:

- Is this method appropriate for what we are trying to find?
- Are there any potential errors that may occur during the process?

By doing so, we can be more confident with our analysis of the research.

Throughout this chapter, we will take a closer look at what has been discussed above, and how we can make sure that the conclusions from both our own research and the research of others is valid.

Developing an investigation

Once we have a clearer understanding of the topic, we are able to further refine the research question and begin to develop the scope of the research investigation, as well as writing the introduction section of our investigation.

Is it dangerous to drink water from a plastic water bottle that's been left in a warm environment?



Introduction

Bisphenol A (BPA), as shown in figure 5, is a monomer used to produce a strong plastic polymer known as polycarbonate *(Vogel, 2009)*. Originally developed for medical uses, BPA is now used to produce strong plastics found in everyday products such as bottles, food containers and pipes. Due to its widespread use, there is growing concern over the health effects of using BPA.

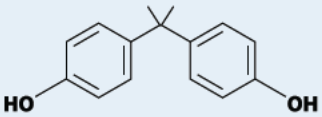


Figure 5 Structure of a bisphenol A (BPA) monomer.

In water-containing products, BPA is known to leach into the water over time *(Le et al, 2008)*. When consumed, this can result in the disruption of hormone-driven processes in the body *(Diamanti-Kandarakis et al. 2009)*. This is potentially dangerous as there are many reactions in the body that are driven by hormones.

In order to be released into the water, BPA needs to be hydrolysed. Due to this, factors such as heat, pH conditions and microwaves can all affect the extent to which BPA is released into its surrounding environment.

Aim

This experiment aims to test the effect of heat on the release of BPA into water.

Hypothesis

It's expected that if the breakdown of BPA from plastic depends on the heat of the surrounding environment, then an increase in temperature would result in an increase in the concentration of BPA found in the water.

As we can see based on our example experiment, it is important to ensure that our research is based on and is linked to theoretical concepts. Although these links are first introduced here, the chemical ideas need to be weaved throughout the whole research investigation. Understanding the underlying scientific theory not only helps us to develop an aim and hypothesis for the investigation, it also helps us to explain the results of the experiment.

While the **aim** of an experiment is a statement about the purpose of the research, the **hypothesis** describes what we expect to happen as a result of the experiment. To develop the hypothesis we first need to understand the main variables in the experiment. More specifically, we need to be able to identify the variable that is being manipulated by the experimenter, the **independent variable**, and the variable that is affected by the change, the **dependent variable**. We can then develop a prediction called the hypothesis that proposes how the independent variable will affect the dependent variable. All variables apart from the independent variable are kept consistent throughout the whole experiment and are known as **controlled variables**.

The simplest way to develop a hypothesis can be seen in table 1.

Table 1 Framework for developing a hypothesis. IV stands for independent variable and DV stands for dependent variable.

	Phrase outlining relationships between DV and IV		Trend indicating effect on the DV		Trend indicating action of the IV
If ... the DV ...	... depends on ... ... results from ... ... is affected by ... ... is related to ...	**... then ...**	... show an increase/decrease ... ... be greater than/less than ... ... be larger/smaller ...	**... when ...**	... increased/decreased ... ... greater/less ... ... large/small ...

For our experiment involving BPA in water bottles, a hypothesis could be:

'If the level of BPA in water is affected by exposure to heat, then an increase in temperature will result in an increase in the level of BPA found in the water.'



Theory summary

- Scientific ideas are ideas that have been developed through experimentation following the scientific method.
- Non-scientific ideas are ideas that have not been developed as a result of the scientific method.
- Research needs to be conducted prior to an investigation to identify current knowledge within the field of interest.
- The aim and hypothesis identify the purpose and prediction of the research respectively.
- Variables are elements of an experiment that can be changed, controlled or measured.

1A QUESTIONS

Theory review questions

Question 1

Scientific ideas are purely developed through past experiences.

A True

B False

Question 2

Scientific ideas only need to be proven once.

A True

B False

Question 3

'Art is in the eye of the beholder, and everyone will have their own interpretation', E.A. Bucchianeri.

Art interpretations can be considered to be

A scientific ideas.

B non-scientific ideas.

Question 4

I	II	III	IV
Atoms are the smallest unit of matter.....	It can be concluded that atoms have a mass in the centre of their structure	Based on the results, it can be seen that...	1. Set up the cathode ray 2. Place two charged plates around the cathode ray

The order which the above tabs should be placed in order to satisfy the scientific method is

A I, III, IV, II.

B I, IV, III, II.

Question 5

Scientific research is built upon previous ideas.

A True

B False

Question 6

All conclusions from research investigations are completely accurate.

A True

B False

Question 7

Textbooks are a

A primary source.

B secondary source.

Question 8

The results of every research investigation support the current theory within the topic of interest.

A True

B False

Question 9

The hypothesis can help us to determine the

A controlled variables.

B independent variable.

Question 10

The observations made during an experiment are mainly due to the

A controlled variables.

B independent variable.

Deconstructed questions

Use the following information to answer questions 11–13

The composition of egg shells is very similar to the enamel found on teeth. Col wanted to conduct an experiment to identify factors that contribute to the breakdown of tooth enamel. Below is a short excerpt of the experimental write up:

Materials:

- Egg shells
- Cola
- Lemon juice
- Milk
- Timer
- 4 × Beakers

Method:

1. Label four beakers from A to D.
2. Place a 1 cm piece of egg shell in each beaker.
3. In beaker A, add 20 mL of cola.
4. In beaker B, add 20 mL of lemon juice.
5. In beaker C, add 20 mL of milk.
6. In beaker D, add 20 mL of water.
7. Leave experiment for 2 hours.
8. Record results.

Question 11 (1 MARK)

The main difference between beakers A to D is

A the volume of substance added.
B the type of liquid used.
C the amount of egg shell used.
D the length of time the experiment.

Question 12 (1 MARK)

The variable(s) kept constant in this experiment is/are:

I The volume of substance added
II The type of liquid used
III The amount of egg shell used
IV The duration of the experiment

A I only
B I & II only
C I & III only
D I, III & IV only

Question 13 (1 MARK)

Identify the dependent variable in the experiment.

Exam-style questions

Within lesson

Question 14 (2 MARKS)

It's not quite known where Zodiac signs originated from, however, it is believed that people used to refer to the movement of the sun and stars to predict the future. This eventually led to the development of the zodiac signs, each of which was based on the interaction between constellations and the placement between the sky and the sun during certain periods.
Nowadays, many people look to these signs to help them explain current and/or future events. What type of idea would this be considered as? Explain.

Question 15 (3 MARKS)

Bug zappers use visible light to attract bugs. Edwin wanted to determine which wavelength of the visible light spectrum would be most effective in attracting bugs. The wavelengths of each colour in the visible light spectrum are shown.

a Identify the independent variable for this experiment. (1 MARK)

b Write an appropriate hypothesis for this experiment. (2 MARKS)

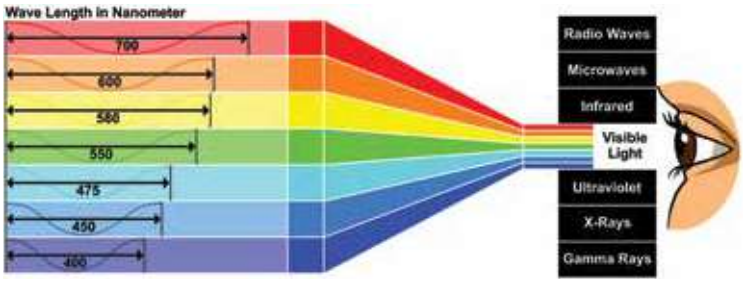


Image: udiax/Shutterstock.com

Exam-style question hints

14: Ideas do not need to be tested in order to be believed. **15a:** The independent variable is deliberately changed in an experiment. **15b:** The hypothesis is a description of the potential observations in an experiment.

1B CONDUCTING AN EXPERIMENT

In this lesson, we will be learning about the factors that we need to take into consideration when selecting and conducting an experiment.

1A Scientific research	1B Conducting an experiment	1C Interpreting data	1D Experimental factors affecting data	1E Writing scientific material

Study design dot points

- the characteristics of laboratory techniques of primary qualitative and quantitative data collection relevant to the investigation: sampling protocols; gravimetric analysis, acid-base titrations and/or pH measurement; precision, accuracy, reliability and validity of data; and minimisation of experimental bias
- ethics of and concerns with research including identification and application of relevant health and safety guidelines
- the influence of social, economic, environmental and ethical factors relevant to the selected chemical investigation

Key knowledge units

Qualitative data collection	2.3.2.1
Quantitative data collection	2.3.2.2
Social, environmental and economic influences	1.3.5.1
Ethical considerations	1.3.5.2 & 2.3.3.1
Safety guidelines	2.3.3.2

Key terms and definitions

Primary data original data collected firsthand by researchers

Secondary data data that has been previously collected that is now accessible to different researchers

Qualitative data non-numerical (descriptive) data collected based on observations taken during an experiment

Qualitative analysis technique that determines the composition of a sample

Quantitative data numerical data collected during experiments

Quantitative analysis technique that identifies the amount of substance present

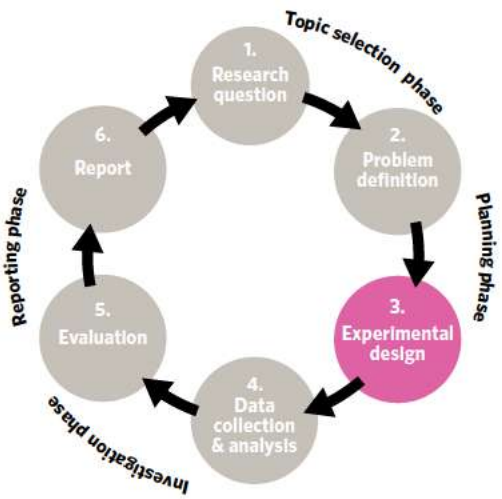
Social considerations taking into consideration the effect on others

Economic considerations taking into account the associated costs (e.g. time and money)

Environmental considerations taking into account the environmental implications

Ethical considerations taking into account the effect on other living organisms (e.g. humans and animals)

Material Safety Data Sheet (MSDS) document that outlines the health and safety information associated with different materials and chemicals



Qualitative data collection 2.3.2.1



Primary data – 'first hand' data

Secondary data – 'second hand' data

OVERVIEW

Some experiments give observational data.

THEORY DETAILS

One of the main ways in which we can continue to develop scientific theory is through conducting experiments to obtain data, which is used to develop conclusions about the area of study. The data that is evaluated can either be collected by the original experimenters, also known as **primary data**, or can be collected indirectly from other experiments which is referred to as **secondary data**.

Another way in which we can differentiate data is by the information that is provided by the data. Different experiments are able to collect different types of data which means that researchers can select the most appropriate experimental and data collection method for their topic of investigation.

During the earlier years in high school, a lot of the data obtained from experiments was based on observations made during reactions. Colour changes, smells and temperature changes are all examples of the types of data that may have been collected. This type of data is classified as **qualitative data**. Qualitative data is based on observations and often describes (in words) what has occurred as a result of a reaction. **Qualitative analysis** allows us to gain information about what substances are made of. An example of a technique that can gather qualitative data that we will be studying later in the course is acid-base titrations. One component of acid-base titrations involves observing a colour change, which signifies the point at which the experiment should stop. This type of data is an example of qualitative data. The advantage of collecting this type of data is that it can give a quick indication as to whether or not a reaction has taken place. Table 1 shows examples of qualitative data collected from different reactions.

Table 1 Qualitative data collected from different techniques/reactions.

Reaction	Example of qualitative data collected
Acid-base titrations	Colour change
Precipitation reactions	Formation of a solid
Esterification	Fruit odour

Quantitative data collection 2.3.2.2

OVERVIEW

Some experiments give numerical data.

THEORY DETAILS

Unlike qualitative data, **quantitative data** is expressed in numbers and can be used to calculate the amount of a substance. This type of data is important as it can help us determine the amount of a substance that is needed for a desired reaction, and also helps us to predict the product that can be formed. Quantitative data is obtained through **quantitative analysis** and examples of quantitative analytical techniques that we will study later in the course include gravimetric analysis, acid-base titrations and HPLC. These are shown in table 2.

Table 2 Quantitative data collected from different techniques.

Technique	Example of quantitative data collected
Gravimetric analysis	Mass of precipitate formed
Acid-base titrations	Concentration of unknown sample
HPLC	Retention time of different components

When selecting appropriate methodology for an experiment, it is important to understand the type of data that is required to answer the research question. If we pick methodology that is inappropriate for the investigation, the results may not address the research question which could jeopardise the validity of the experiment. We will learn about the concept of validity in greater depth in the following lessons.

For example, if we were trying to investigate the relationship between an individual's height and the speed at which they can run 100 m, it would be inappropriate to measure the temperature change over time as it does not address the research question. As such, the results obtained would not be able to be analysed.

Not only is this important for our own research, this is also the type of thinking we need to adopt when reviewing other research.

For our experiment, we are trying to collect information about the levels of BPA in water. In particular, we want to collect quantitative data so that we will be able to compare BPA levels as a result of different conditions. Based on the information gathered during the research phase, we will be using HPLC as our analytical technique to collect the required quantitative data. In the experiment, we want to measure the BPA levels found in the bottles of water exposed to different temperatures over a two hour period.



Social, environmental and economic influences 1.3.5.1

OVERVIEW

Scientific research has greater implications outside the scientific community.

THEORY DETAILS

Regardless of the type of experiment conducted, when designing a research investigation, we need to make sure that we consider the wider implications of the study. Questions that could be asked during this process include:

- Who will be impacted by the results of the study?
- How will the results of the study contribute to the decision-making process at a larger scale?
- What are the implications of the study with respect to the environment?
- What are the potential environmental effects of the substances used during the experiment?

Although it may seem like the experiment we conduct may not have a large impact on the wider community, the decisions that are made today are a result of the mountain of work that has been done by scientific researchers. For example, the different strategies and agreements that have been developed to reduce our impact on the environment come from the countless studies that have been conducted in all fields of science. When conducting research, we need to think about the:

- **Social considerations** – who is affected by the research?
- **Economic considerations** – what are the associated money and resourcing costs?
- **Environmental considerations** – how is the natural environment affected by the research?

In terms of our BPA-related experiment, the results of the research may influence consumer behaviour when it comes to the use and storage of BPA plastic bottles. Thinking beyond the potential effect on the immediate user, depending on the results obtained through the experiment, this could have a wider implication on other ecosystems if BPA bottles are found as a pollutant in waterways as a result of littering.



The results from scientific research have flow-on effects that can impact many different people and industries in different ways. Figure 1 shows an example of the potential effects of scientific research around Dichlorodiphenyltrichloroethane (DDT), a chemical that was used as an insecticide.

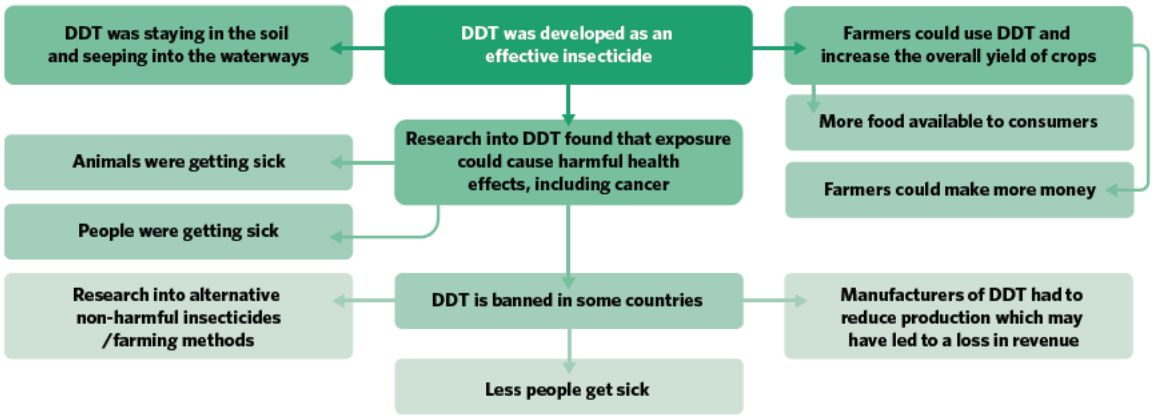


Figure 1 Illustration of the potential impacts of scientific research.

The broader implications of our research study on plastic bottles containing BPA may affect the overall production of these types of bottles, which could potentially have a financial impact on the companies responsible for the production of these materials. At the same time, this could lead to further research into alternative materials and encourage a behaviour shift that can have a positive impact on the whole environment.



Ethical considerations 1.3.5.2 & 2.3.3.1

OVERVIEW

Ethics should always be taken into account when conducting scientific research.

THEORY DETAILS

Conducting scientific research is about more than just scientific theory. Due to the potential implications of the work, we are all responsible for the appropriate conduct and use of the knowledge obtained. As such, we need to make sure that we adhere to **ethical considerations** when conducting research. In other words, we need to make sure that we're trying our best to do the 'right' thing from a moral and principled perspective.

When thinking about all of the ethical issues that may arise from Chemistry-related studies, we can categorise them into two main areas; the internal and external domains. The internal domain consists mainly of the issues related to the experiment itself whereas the external domain consists of the wider societal implications of the research. Table 3 shows examples of the two types of ethical domains:

Table 3 Examples of ethical issues in internal and external domains.

Internal domain	External domain
• Conducting experiments safely • Plagiarism • Data falsification • Acknowledgement of previous research (i.e. citations) • Animal/human testing	• Development of new substances • Environmental impacts • Societal impacts

Due to the importance of ethics in chemistry research, scientists met to develop a document known as the Hague Ethical Guidelines. This guideline outlines the main principles that we should try to abide by when conducting chemical research. This document is not directly related to the VCAA study design however is a good resource that gives further information about ethics in chemistry.

Safety guidelines 2.3.3.2

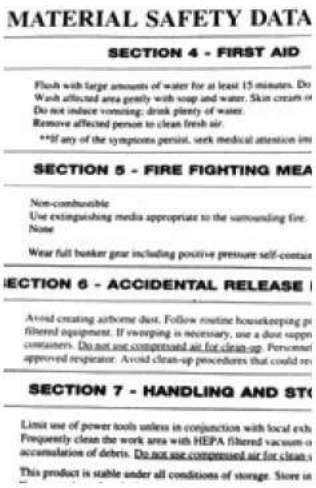
OVERVIEW

All measures should be taken to ensure the safety of everyone involved in scientific investigations.

THEORY DETAILS

All scientific experiments come with their own safety risks. There are many different types of risks, which vary depending on the type of experiment conducted. These risks can arise through the use of certain substances as well as from the experimental procedure itself.

Prior to conducting any experiment, it is important that the experimenter refers to the **Material Safety Data Sheet (MSDS)** of all of the substances that will be used as part of the experiment. The MSDS contains all of the health and safety information relating to materials/substances that will be used, including strategies and guidelines to help mitigate the effects of each. Figure 2 shows sections that are often included in an MSDS.


MATERIAL SAFETY DATA

SECTION 4 - FIRST AID

SECTION 5 - FIRE FIGHTING MEA

ECTION 6 - ACCIDENTAL RELEASE

SECTION 7 - HANDLING AND ST

Figure 2 Sample MSDS of a chemical substance.

As part of the experimental process, scientists use MSDS to develop a risk assessment for the experiment to analyse the potential risks and outline specific strategies to minimize these risks. For example, experiments that result in the production of a potentially harmful gas would be required to be conducted in a fume hood. Other examples of safety measures that can be taken during an experiment include:

- Wearing personal protective equipment eg. lab coat, safety glasses etc.
- Knowing where and how to use the safety showers.
- Proper disposal of chemical waste.

For our BPA-related experiment, we need to collate all of the information related to the substances that will be used in the experiment and develop a plan to minimise safety risks. This will mainly involve information about BPA that is used to create standards. This should be included as part of the 'Materials and Methods' section of a scientific investigation.



Theory summary

- Qualitative data is mainly observational.
- Quantitative data gives us numerical information about substances.
- Scientific findings have social, environmental and economic implications.
- There are many ethical concerns related to chemical investigations that should be considered by the people conducting the research.
- It is very important to make sure that experimenters are well informed about the nature of the experiment they are conducting so that they can put strategies in place to minimise the potential risks.

1B QUESTIONS

Theory review questions

Question 1

Qualitative analysis is optimal when

A an experimenter requires information about colour differences.

B an experimenter requires information about exact temperature changes.

Question 2

This image shown represents the collection of

A qualitative data.

B quantitative data.



Image: ggw/Shutterstock.com

Question 3

The image in question 2 shows the collection of

A primary data.

B secondary data.

Question 4

Secondary data can be

A quantitative only.

B both quantitative and qualitative.

Question 5

All scientific research findings have positive implications for all areas of society.

A True

B False

Question 6

Scientists are responsible for all misuse of research findings.

A True

B False

Question 7

Scientists who conduct research in the same area are able to use each other's findings without needing to acknowledge the source.

A True

B False

Question 8

When a new substance or scientific technique is developed, it should be shared with the world immediately.

A True

B False

Question 9

Scientists need to conduct a primary test of all the substances they are going to use before performing their actual experiment in order to find out the potential risks of using the substances.

A True

B False

Question 10

Safety considerations for experiments only need to be documented after the experiment has been conducted, and only if issues arise during the experiments.

A True

B False

Deconstructed questions

Use the following information to answer questions 11–13

When increasing atmospheric concentrations of carbon dioxide (CO_2) are discussed in the context of climate change, a silver lining is often postulated: 'CO_2 is plant food.' That is, higher concentrations of CO_2 are generally acknowledged to stimulate plant photosynthesis and growth, with potential benefits for the productivity of the cereal crops that remain the world's most important sources of food [1]. Cereal crops feed not just humans but also the animals that are important sources of protein for many. Since the mid-1960s, cereal production increased by approximately a billion tons [2], yet accelerated progress in agriculture is needed to keep pace with population growth anticipated to reach between 9–10 billion by 2050 [3], and to achieve Sustainable Development Goal (SDG) 2 to end hunger, achieve food security and improved nutrition, and promote sustainable agriculture.

But food security is about more than just production. While increases in CO_2 may make some crops grow more quickly, research shows that higher CO_2 concentrations can also reduce the nutritional quality of staple crops, from potatoes to barley, rice to wheat [4–6].

Source: Ebi, K. L., & Ziska, L. H. (2018). Increases in atmospheric carbon dioxide: Anticipated negative effects on food quality. PLoS medicine, 15(7), e1002600.

Question 11 (1 MARK)

The effect of the CO_2 increase can be experienced by

A direct consumers.

B farmers.

C food manufacturers.

D all of the above.

Question 12 (1 MARK)

What effect would an increase in CO_2 have on the health of consumers?

A Consumers will be able to eat less to obtain the same amount of nutrients.

B Consumers will have to eat more to obtain the same amount of nutrients.

C Consumers will spend less on food.

D Consumers will eat less cereal.

Question 13 (2 MARKS)

Describe an environmental issue evident in the passage.

Exam-style questions

Within lesson

Question 14 (2 MARKS)

Cecelia and her team were conducting experiments to develop a new type of drug targeted at supporting patients with diabetes. As part of their work, they conducted extensive research on diabetes and the current available therapies and came across a study that was similar to theirs. The team used the results of the study to direct their research.

a What type of data, primary or secondary, was Cecelia and her team collecting to do this research? (1 MARK)

b In an effort to finish the research on time, Cecelia copied and pasted the research conducted by the other team and used it as her own. What is the main ethical issue around what she did? (1 MARK)

Question 15 (4 MARKS)

Luis was given two unknown samples of metallic substances and asked to determine the name of each. In his experience, he knew that he could use the flame test to differentiate between the two metallic substances.

Upon conducting the experiment, Luis noticed that each sample gave off a different coloured flame.

a Identify whether this type of analysis is qualitative or quantitative. (1 MARK)

Luis wanted to use the results of the test to find how much of each substance was present.

b Is it possible for Luis to calculate the amount of substance present based on his results? Explain. (3 MARKS)

Question 16 (5 MARKS)

One of the methods currently used to develop vaccines involves scientists intentionally infecting individuals with viruses, bacteria or parasites. Individuals can be given a vaccine against a particular disease and exposed to the pathogen responsible for causing the disease. In doing so, scientists are able to determine whether or not the vaccine is effective by measuring the extent to which the person shows symptoms of being ill. In comparison to traditional methods of vaccine development, this method is much quicker and requires only a few rounds of testing.

Alternatively, scientists can spend years developing vaccines that are rigorously tested in other contexts before being trialled on humans.

a Identify the main concern discussed in the passage above. (1 MARK)

b Identify one reason why scientists would choose to conduct experiments on humans. (1 MARK)

c Discuss the major disadvantages of both methods of vaccine development. (3 MARKS)

Question 17 (4 MARKS)

Elephant toothpaste can be made using reactants such as hydrogen peroxide and yeast.

A section from the MSDS for hydrogen peroxide is shown.

Classification of the substance or mixture – GHS-US classification		
Oxidizing liquids Category 1	H271	May cause fire or explosion; strong oxidizer
Acute toxicity (oral) Category 4	H302	Harmful if swallowed
Acute toxicity (inhalation) Category 4	H332	Harmful if inhaled
Skin corrosion/irritation Category 1A	H314	Causes severe skin burns and eye damage

Source: www.labchem.com/tools/msds/msds/LC15430.pdf

a Identify two pieces of personal protective equipment that would allow you to safely conduct this experiment. (2 MARKS)

b As part of the preparation for this experiment, a student asked the lab technician, who is not involved in the implementation of the method, to develop a risk assessment. Is the lab technician the appropriate person to develop the risk assessment? Explain. (2 MARKS)

Multiple lessons

Question 18 (5 MARKS)

The following is an excerpt from a research article:

Background

Dietary sugar, especially in liquid form, increases risk of dental cavaties, adiposity, and type 2 diabetes. The United Kingdom Soft Drinks Industry Levy (SDIL) was announced in March 2016 and implemented in April 2018 and charges manufacturers and importers at £0.24 per litre for drinks with over 8 g sugar per 100 mL (high levy category), £0.18 per litre for drinks with 5 to 8 g sugar per 100 mL (low levy category), and no charge for drinks with less than 5 g sugar per 100 mL (no levy category). Fruit juices and milk-based drinks are exempt. We measured the impact of the SDIL on price, product size, number of soft drinks on the marketplace, and the proportion of drinks over the lower levy threshold of 5 g sugar per 100 mL.

Source: Scarborough, P., Adhikari, V., Harrington, R. A., Elhussein, A., Briggs, A., Rayner, M., Adams, J., Cummins, S., Penney, T., & White, M. (2020). Impact of the announcement and implementation of the UK Soft Drinks Industry Levy on sugar content, price, product size and number of available soft drinks in the UK, 2015–19: A controlled interrupted time series analysis. PLoS medicine, 17(2), e1003025. https://doi.org/10.1371/journal.pmed.1003025

a Based on the information, write a hypothesis for the proportion of drinks with over 5 g sugar per 100 mL as a result of the introduction of SDIL. (1 MARK)

b What is the primary type of ethical concern in this experiment? (1 MARK)

c Describe a potential implication of the SDIL and how it would affect society. (3 MARKS)

Key science skills

Question 19 (4 MARKS)

Jo wanted to determine the concentration of different minerals in bottled mineral water. As part of the investigation, three different brands of mineral water were used.

a Identify the independent variable. (1 MARK)

b List three variables that should be kept constant throughout the experiment. (3 MARKS)

Exam-style question hints

14a: The data type depends on the source of the data. **14b:** Scientific research reports should be written in the experimenter(s) own words **15a** There are two main data types based on what the data measures. **15b:** To calculate the amount of substance, numerical data needs to be collected. **16a:** Special consideration needs to be taken into account when experimenting on humans. **16b:** It is important to make medicine available in a timely manner. **16c:** The disadvantages of each method depend on how we analyse the experiment. **17a:** Safety precautions for experiments are selected based on the risks relating to substances used. **17b:** As part of the planning process, researchers are required to analyse potential risks of the experiment. **18a** Generating a hypothesis requires us to understand the purpose of the experiment. **18b:** In general, ethical concerns relate to the experiment or the implications of the research itself. **18c** Taxes are a way in which the government can help to shift behaviour that is more beneficial for society. **19a** Independent variables are changed by the experimenter. **19b:** All conditions apart from the independent variable should be kept constant in an experiment.

1C INTERPRETING DATA

In this lesson, we will learn about the different ways in which we can interpret and represent data obtained from an experiment.

1A Scientific research | 1B Conducting an experiment | 1C Interpreting data | 1D Experimental factors affecting data | 1E Writing scientific material

Study design dot points

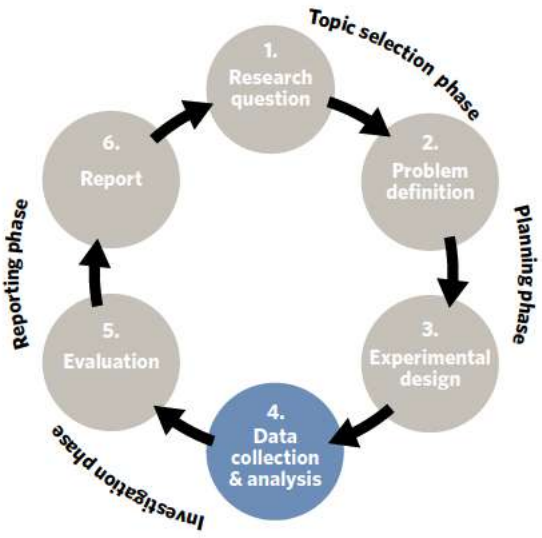
- methods of organising, analysing and evaluating primary data to identify patterns and relationships including identification of sources of error and uncertainty, and of limitations of data and methodologies
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations
- observations and experiments that are consistent with, or challenge, current chemical models or theories
- the nature of evidence that supports or refutes a hypothesis, model or theory
- the key findings of the selected investigation and their relationship to solubility, concentration, acid/base and/ or redox concepts

Key knowledge units

Identifying patterns in data	2.3.4.1 & 1.3.3.1
Drawing conclusions from data	2.3.5.1 & 2.3.6.1
Linking chemical concepts to conclusions	2.3.8.1

Key terms and definitions

Control group (in chemistry) sample that is subjected to the same conditions as all other samples without the independent variable



Lesson links

This lesson builds on:

- 1B Conducting an experiment

Data collected from experiments is analysed and used to develop a deeper understanding of the research question.

Identifying patterns in data 2.3.4.1 & 1.3.3.1

OVERVIEW

Experiments generally show a relationship between the variables being studied.

THEORY DETAILS

The main purpose of an experiment is to identify a casual relationship between the independent and dependent variables. Unless there are flaws in the experiment, we will generally be able to see a pattern emerge from the data collected.

The way we present data can make trends easier to spot. For example, it is easier to identify trends in quantitative data when it is represented as a line graph than it would be if the same data was presented as a pie chart. Table 1 shows the common forms of data representation and the type of data they are mostly used for.

Table 1 Characteristics of common data representation formats.

Type of data representation	Example	Type of data being represented	Advantage
Line graph	Calibration curves	Quantitative	Good for displaying trends over time; also helpful when predicting trends over time
Bar graph	Water sampling in different areas	Both qualitative and quantitative	Easy to visualise and compare different categories of data
Pie chart	Composition of fuel	Both qualitative and quantitative	Used to show proportions (e.g. percentage)
Table	Gravimetric analysis	Both qualitative and quantitative	Helps to organise and compare multiple sets of values

When constructing and analysing graphs, it's important to know the variable that is being represented on each axis. As seen in figure 1, the y-axis represents the dependent variable, whereas the x-axis represents the independent variable.

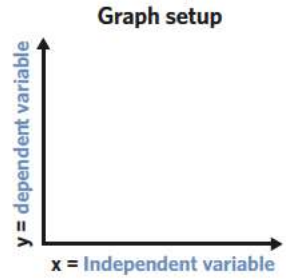


Figure 1 Representation of data in graphs.

When interpreting graphs, we need to think about how the independent variable affects the dependent variable. For example, consider the graph shown in figure 2.

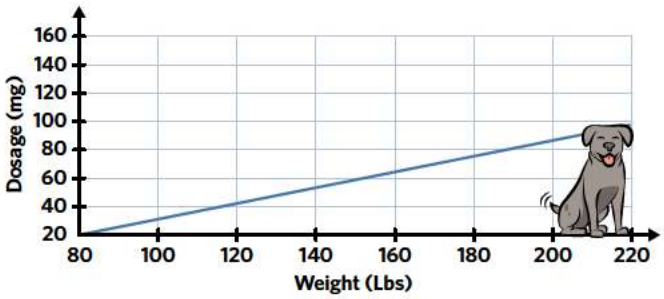


Figure 2 Graph showing information about the dosage of medicine for dogs.

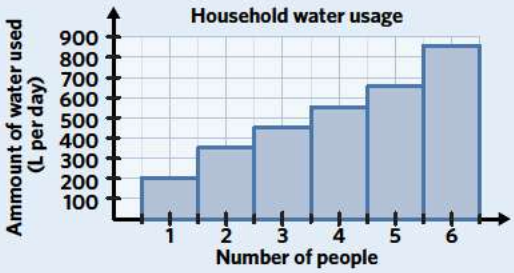
In the graph, we can see that the dosage of medicine given to a dog (vertical axis) increases as the weight of the dog (horizontal axis) increases.

Regardless of whether we are interpreting our own data collected from our experiment, or we are interpreting experimental data from an experiment in a journal article, we are using the same strategy to identify trends. Being able to do so will help us determine the accuracy of the claims made by experimenters about their own research.

Worked example 1

Water companies use data to estimate the water usage of different sized households. The graph provided shows the amount of water usage based on the number of people in the house per day.

Predict the amount of water used in a household per day containing seven people.



What information is presented in the question?

A bar graph representing water usage.

What is the question asking us to do?

Predict the amount of water used in a house with seven people.

What strategy(ies) do we need in order to answer the question?

1. Identify the general trend in the data presented in the graph.
2. Find the average change in the amount of water used as the number of people increases.
3. Determine a possible range for the water usage per day for seven people.

Answer

Overall, it can be seen that the more people there are in the household, the more water is used per day.

Upon closer analysis of the increase in water usage per person added, it can be seen that, on average, there is approximately a 130 L increase with each additional person.

Therefore, it is predicted that in a house of seven people, the water usage could range from approximately 860–990 L.



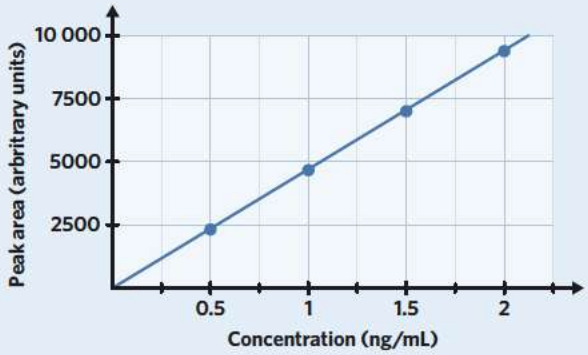
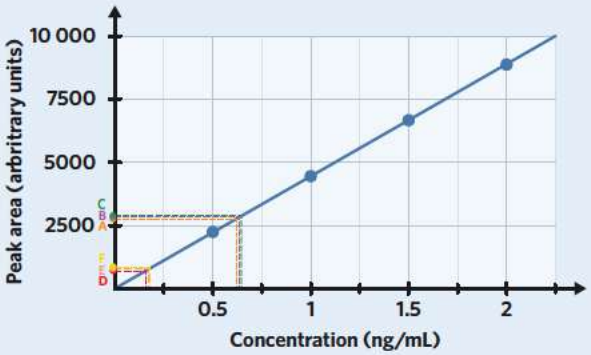


Figure 3 Calibration curve of known standards of BPA.

The graph provided shows the calibration curve developed from the use of BPA standards. As we can see by the trend shown, as the concentration of BPA increases, the area under the curve also increases. We can use this graph to identify the concentration of BPA in the different samples in our experiment.

We can cross reference these values with the calibration curve to determine the concentration of BPA present in each sample.



Sample
A-BPA bottle at 4°C
B-BPA bottle at 20°C
C-BPA bottle at 30°C
D-BPA-free bottle at 4°C
E-BPA-free bottle at 20°C
F-BPA-free bottle at 30°C

Figure 4 Determination of the concentration of BPA in samples A–F.

Based on the information given, it can be seen that as the bottle is left in higher temperatures, there is an increase in the concentration of BPA in the water.

Drawing conclusions from data 2.3.5.1 & 2.3.6.1

OVERVIEW

The trends represented by the data can help us develop a better understanding of the scientific theory related to the research.

THEORY DETAILS

The data collected through scientific research can help address the nature of the hypothesis. The results of an experiment do not have to necessarily support the hypothesis; in fact, the greatest learnings come from experiments where the data refutes the hypothesis.

Firstly, to know that the trend in the experimental data collected was due to the independent variable, we need to make sure that a **control group** is present as a point of comparison to show that any change in the experiment was due to the impact of the independent variable. A control in an experiment is set up in a way that doesn't allow the independent variable to impact the results (i.e. the independent variable isn't applied or changed in the control sample). This is important as it allows us to attribute the changes in the experiment to the independent variable, and the effect of other variables.



Back to our experiment. Of the six samples, samples A–C involve BPA-containing bottles whereas D–F do not. The purpose of D–F is to act as a control to show that the changes in BPA concentration were due to the BPA present in the bottles that were subjected to different conditions, instead of just the change in temperature. To ensure the quality of our experiment, all other variables, such as volume of water and duration of the experiment, were controlled (and therefore were our controlled variables). Based on the result above, we can see that the data supports the initial hypothesis made.

One point to remember is that the trends seen in the data collected in an experiment do not necessarily illustrate a causal relationship; that is, it does not mean that the changes seen in the dependent variable are solely due to the independent variable. There may be a correlation, which is that the changes observed in the dependent variable are associated with the independent variable, but it is possible that other variables could also be responsible for this change. As such, we can't say that the result was caused by the independent variable; we can't establish causation. For example, based on the data presented in figure 5, it can be seen that as the cost of a packet of chips increases, the amount of sour cream consumed also increases. However, this does not necessarily mean the cost of chips is the reason why people eat more sour cream.

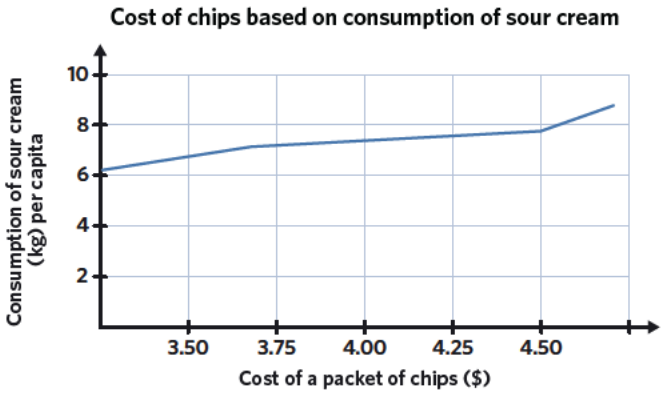


Figure 5 Line graph representing the relationship between sour cream consumption and cost of chip packets.

Furthermore, there may be factors that affect the validity and reliability of our conclusions; a concept that we will explore in the next lesson.



For our water bottle experiment, we need to unpack the data to understand the implications of our observations. This will be the bulk of the discussion section of our experimental report.

Discussion

For samples A–C, it can be seen that an increase in temperature is correlated with an increase in the concentration of BPA measured. The samples placed in the 30°C and 20°C environments and in the fridge measured BPA concentrations in ng/mL of 0.63, 0.61 and 0.58 respectively.

BPA can be released into the contents of water bottles when the plastic begins to degrade. As shown in previous studies, an increase in temperature results in an increase in the degradation of plastic (Biedermann-Brem & Grob., 2008). In particular, an increase in temperature results in the increase in the rate at which the plastics are degraded. In the context of BPA-containing plastic water bottles, this could lead to an increase in the rate at which BPA is released into the water. This trend is shown in the results of this experiment. As can be seen in the experimental results, there was a greater concentration of BPA found in the sample subjected to the highest temperature (sample C). It is possible then that at higher temperatures, there could be a greater release of BPA from plastic, resulting in a higher concentration of BPA recorded.

Linking chemical concepts to conclusions 2.3.8.1

OVERVIEW

The conclusions developed from research need to be explained by scientific theory.

THEORY DETAILS

When evaluating research the conclusions drawn need to be supported both by the experimental data collected during the experiment as well as existing theory. Regardless of whether or not the hypothesis is supported, it is important that we recognise the scientific theory that could explain the observations made.

In our water bottle experiment, the main ideas related to the experiment involve:

- Concentration
- Bonding
- Temperature and chemical reactions

As a result, discussions about and explanations of the data need to be based on our current understanding of these key ideas. This thought process also applies when we are evaluating other research.

Theory summary

- Results generally show a relationship between the independent and dependent variables.
- A control group is important to show that the changes were due to the variable being tested.
- The way in which we represent data depends on the type of data collected.
- Conclusions generated from data need to be supported by scientific theory.

1C QUESTIONS

Theory review questions

Question 1

Experimental data is always shown as a graph.

A True

B False

Question 2

The type of data collected helps to determine the

A way in which data is presented.

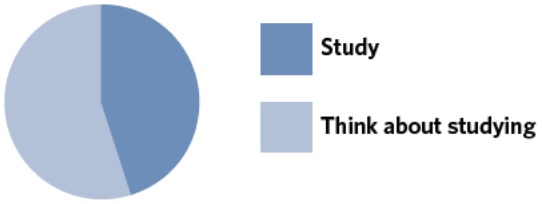
B variables in the experiment.

Question 3

The figure provided would be best used to represent

A qualitative data.

B quantitative data.



Question 4

Generally speaking, experimental data shows

A trends between the controlled and independent variables.

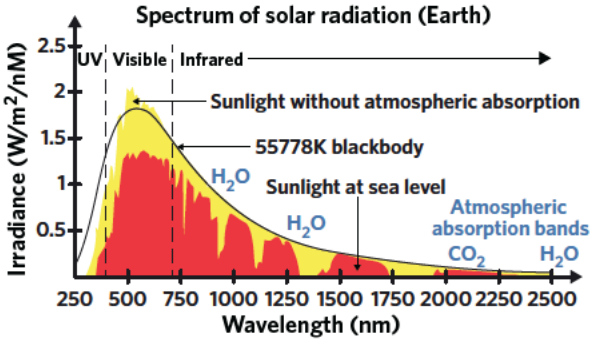
B trends between the independent and dependent variables.

Question 5

In the graph provided, wavelength represents the

A independent variable.

B dependent variable.



Question 6

In the graph provided in question 5, the _____ is affected by the _____.

A wavelength, irradiance

B irradiance, wavelength

Question 7

For an experiment measuring the effect of the amount of water used on the growth of a sunflower seed, a suitable control group would be

A a sunflower seed receiving twice the amount of water compared to all of the other sunflower seeds.

B a sunflower seed receiving no water.

Question 8

The following statement was made about an experiment:

'As the temperature increases, it can be seen that students complete more homework.'

In this experiment, it is most likely that

A the change in temperature causes students to do more homework.

B the change in temperature is related to students completing homework.

Question 9

It is sometimes difficult to prove causation because

A results are always affected by the same variable.

B results can be affected by multiple factors.

Question 10

Conclusions in research will always prove a hypothesis correct.

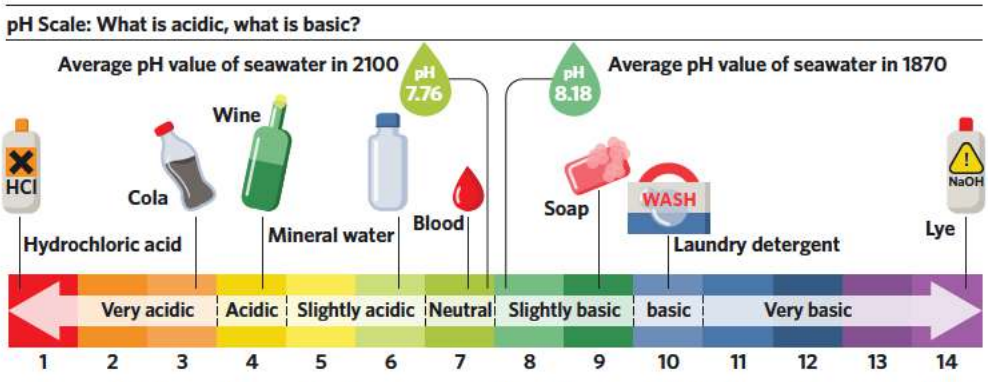
A True

B False

Deconstructed questions

Use the following information to answer questions 11–13

Scientists use pH to measure how acidic or basic substances are. For example, laundry detergents are considered basic whereas wine is considered acidic.



The difference may seem small, but the decline in pH value frm 1870 to 2100 would mean a 170% increase in acidity Much smaller changes already pose problems for many sea creatures.

Question 11 (1 MARK)

Substances that are neutral have a pH of approximately

A 5.

B 6.

C 7.

D 8.

Question 12 (1 MARK)

Generally speaking, when cola is added to water, the solution would become

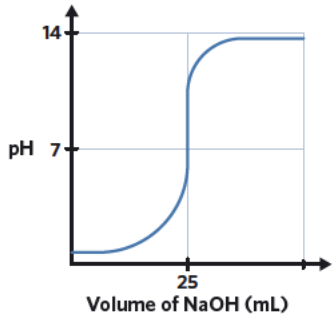
A more acidic.
B more basic.
C less acidic.
D neutral.

Question 13 (2 MARKS)

The graph provided shows the pH levels in a solution to which NaOH was added over time.

Would the original solution (before NaOH was added) be considered as acidic or basic? Explain.

Adapted from VCAA 2016 Exam MCQ 6



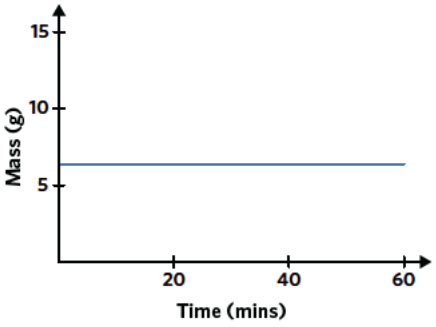
Exam-style questions

Within lesson

Question 14 (3 MARKS)

Our current understanding of science is built from the knowledge developed through a large body of research.

Consider the graph below, which represents data collected from an experiment, where a beaker containing the reaction was left on a mass balance and weighed over time.



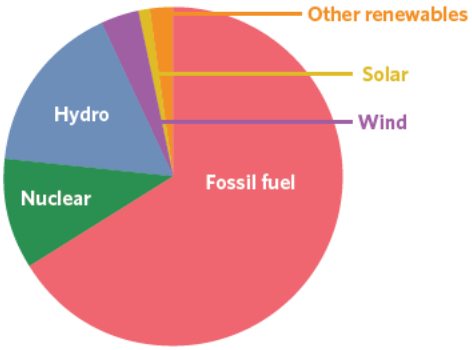
a Identify the trend shown in the graph. (1 MARK)

b Use any of the following chemical concept(s) to justify the conclusion(s) drawn from the graph provided. (2 MARKS)

I The law of conservation of mass states that no mass is ever lost or gained overall.
II Energy is never 'lost', it is merely transformed.
III There are approximately 6.02×10^{23} atoms/molecules/ions in a mol.

Question 15 (4 MARKS)

The world requires different types of energy in order to sustain all of the work that occurs. The graph provided shows the world's electricity sources in 2015.



a Identify the overall breakdown of energy sources shown in the graph. (1 MARK)

b Fossil fuels are formed over millions of years and generally extracted through deep drilling into the earth. The use of fossil fuels produces substances such as carbon dioxide, sulfur dioxide and nitrogen oxides. Given that our population continues to grow, what would you expect to happen over the next 10 years with respect to the usage and impact of fossil fuels? (3 MARKS)

Multiple lessons

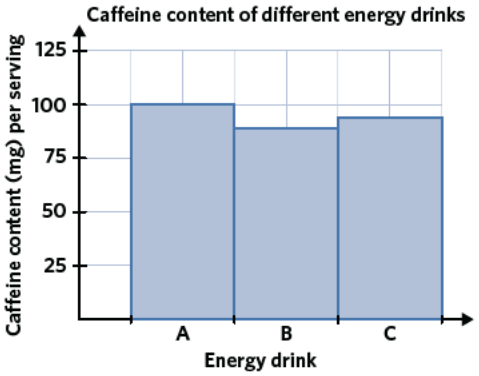
Question 16 (4 MARKS)

Energy drinks that contain a large amount of caffeine are becoming increasingly popular amongst teenage students. On average, each serving contains approximately 100 mg of caffeine, which is over double the amount of caffeine contained in a soft drink.

A student wished to study the caffeine content in various energy drinks.

a What type of data would be collected in the experiment? (1 MARK)

The student presented the data in the graph to the right.



b What is the general information presented in this graph? (1 MARK)

At high levels, caffeine is known to cause symptoms such as:

- Headaches
- Sleeping problems
- Increase in heart rate

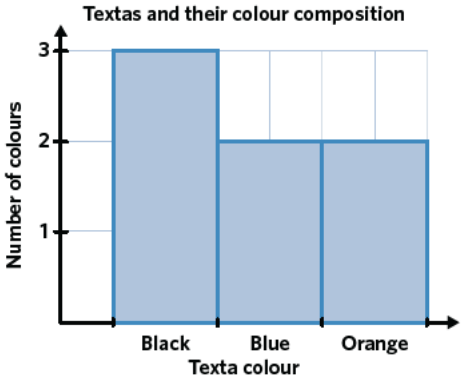
To further investigate the impact of caffeine on adolescents, the student proposed an experiment involving her friends.

c Identify a major ethical concern of this experiment. (1 MARK)

d Given that the student wanted to investigate the impact of caffeine-containing energy drinks on sleep, identify a suitable control for the experiment. (1 MARK)

Question 17 (3 MARKS)

A student wanted to conduct an experiment to determine the composition of different coloured textas. The graph provided shows the results of the experiment.



a Identify the dependent variable. (1 MARK)

b List two variables that need to be kept constant throughout the experiment. (2 MARKS)

Exam-style question hints

14a: Trends in data relate to the response of the dependent variable as a result of manipulating the independent variable. **14b:** Conclusions need to be supported by appropriate theory. **15a:** The information presented by a pie chart depends on the detail shown in the graph. **15b:** Data can be used to predict future outcomes. **16a:** The type of data depends on the nature of the data itself. **16b:** Trends in data relate to the response of the dependent variable as a result of manipulating the independent variable. **16c:** Ethical concerns can affect humans, animals, society and the environment. **16d:** Controls groups are used to demonstrate the effect of the independent variable. **17a:** Each axis on a graph represents a different variable. **17b:** Outside of the independent and dependent variables, every other variable in an experiment should be controlled.

1D EXPERIMENTAL FACTORS AFFECTING DATA

In this lesson, we will learn about the factors that affect the quality of the data collected during an experiment and how this can impact the research findings.

1A Scientific research | 1B Conducting an experiment | 1C Interpreting data | 1D Experimental factors affecting data | 1E Writing scientific material

Study design dot points

- the nature of evidence and information: distinction between weak and strong evidence, and scientific and nonscientific ideas; and validity, reliability and authority of data including sources of possible errors or bias
- the characteristics of laboratory techniques of primary qualitative and quantitative data collection relevant to the investigation: sampling protocols; gravimetric analysis, acid-base titrations and/or pH measurement; precision, accuracy, reliability and validity of data; and minimisation of experimental bias
- the conventions of scientific report writing including chemical terminology and representations, symbols, chemical equations, formulas, units of measurement, significant figures and standard abbreviations
- methods of organising, analysing and evaluating primary data to identify patterns and relationships including identification of sources of error and uncertainty, and of limitations of data and methodologies

Key knowledge units

Sources of error and bias	1.3.4.3 & 2.3.2.5
Precision and accuracy	2.3.2.3
Significant figures in reports	2.3.9.4
Validity and the three R's	1.3.4.1 & 2.3.2.4
Limitations of data	2.3.4.2

Key terms and definitions

Uncertainty level of doubt regarding the data obtained (for VCE purposes, we are only looking at this from a qualitative perspective)

Error difference between the measured value and the true value

Systematic error error in measurement by the same amount in the same direction every time which also includes errors that are inherent in the experiment

Random error error in measurement that differs in amount each time the experiment is conducted and is usually a 'one-off' error

Bias difference between the average of a large set of measurements and the true value

Personal error mistake or misinterpretation of a reading or method by the experimenter

Precision how close measured values are to each other

Accuracy how close measured values are to their true value

Validity whether or not the experiment and its components including the results address the aim and hypothesis of the research

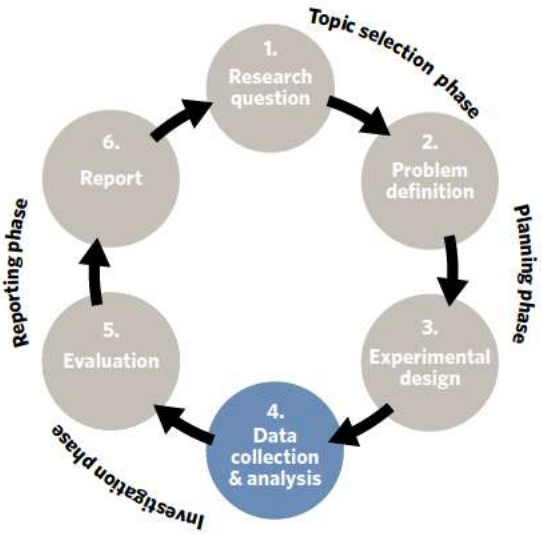
Significant figures number of digits required to express a number to a certain level of accuracy

Lesson links

This lesson builds on:

- 1C Interpreting data

Conclusions drawn from data can be affected by many different factors.



Sources of error and bias 1.3.4.3 & 2.3.2.5

OVERVIEW

Errors can affect the overall results collected.

THEORY DETAILS

When conducting experiments, we always try to make sure that the conditions are well controlled. However, there is still a chance that errors can occur. These errors ultimately affect our results, creating a level of **uncertainty** in the data collected during the experiment. The main sources of error stem from two types of **errors**; **systematic** and **random errors**.

Systematic errors

Systematic errors impact the accuracy of measurements in a consistent manner. The values that are measured are skewed to the same degree; meaning that they are 'off' by the same amount every time. Systematic errors can be built-in to the materials that we use for the experiment or into the methodology. As a result, we can often identify and compensate for these errors.

An example of a common systematic error is in the mass balance. If the mass balance that we use in the laboratory is calibrated in a way that makes the values 0.03 g lower than the true value, then everything that is weighed by the same mass balance will be out by a measure of 0.03 g. Another common example involves the reading of measuring cylinders as shown in figure 1.

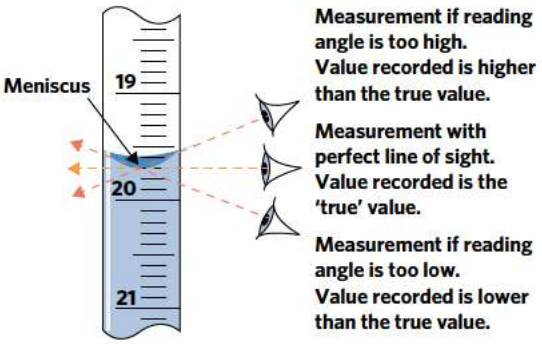


Figure 1 Errors that can arise during the use of measuring cylinders.

When reading the meniscus of a liquid substance, it is important to be positioned in a way that allows the eye line to be directly horizontal to the meniscus; the curve of the liquid. Being above or below the meniscus will result in recordings that are inaccurate; either higher or lower than the true value. These types of errors are known as parallax errors. Unfortunately, these types of errors cannot be improved by repeating the experiment.

To reduce the effect of systematic errors, we should calibrate all apparatus before conducting the experiment and analyse the method to identify steps that could introduce errors. Understanding correct measuring techniques is also beneficial.

For our BPA experiment, a systematic error could occur in the volume of water in all of the water bottles. Since we trusted the labelling of the unopened bottle, we assumed that there was the correct volume of water present in the bottles. A way to rectify this error would be to measure the exact volume of water using scientific apparatus derived from the same brand. Since this could also introduce a parallax error, we would have to be careful to make sure that we use the appropriate technique.

Random errors

These errors are inconsistent and unpredictable. This type of error generally occurs when judgements need to be made quickly about what is occuring in the experiment and/or when an estimation is required. For example, a random error can occur when the volume of a liquid needs to be estimated as a consequence of using a measuring cylinder that is missing lines or numbers. We can minimise this type of error by repeating experiments multiple times to calculate the average of the results.

Systematic and random errors affect an experiment in different ways as shown in figure 2.

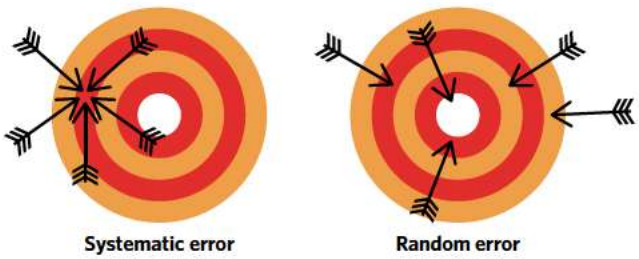


Figure 2 Effect of systematic and random errors on an experiment.

From the perspective of a graph, systematic errors will cause the whole graph to be shifted by the same degree. The trend shown as a result of both errors is generally still the same, however the values will be different to the true value. With respect to random errors, we would have to draw a line of best fit. The effect of both systematic errors and random errors are depicted in figure 3.

We can use the measured values to determine the difference between the average of repeated experiments and the data (inclusive of systematic errors), which is referred to as **bias**. The best way to minimise bias would be to minimize both types of errors in an experiment.

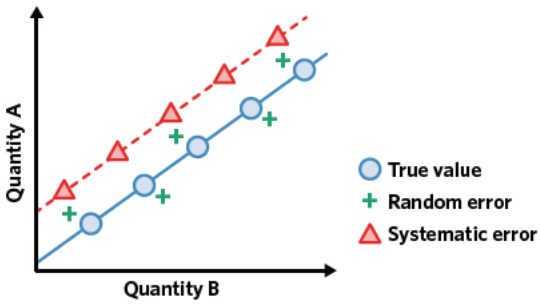


Figure 3 Graphical representation of the effect of systematic and random errors.

Personal errors

Personal errors include mistakes made due to a misinterpretation of information or methodology. These can occur when incorrect formulas are used, collected data is unrelated to the actual experiment or even when the scale of a measuring instrument is misread. Fortunately, these errors can be rectified by repeating the experiment properly.

In regards to our water bottle BPA experiment, a personal error that could occur may involve the extraction of a sample of pure BPA using a pipette that measures mL instead of µL for one of the samples.



Detecting errors in research

Errors can have a big impact on the quality of research. Therefore, it is important that we are able to identify errors in research that we read. This can range from theoretical misinterpretations, to analytical and measurement errors as mentioned above. With respect to errors identified above, some questions that will be useful in identifying errors could be:

- Are there outliers in the data? If so, why?
- Were the measuring instruments calibrated?
- Is the methodology prone to misinterpretation? (i.e. does the methodology outline the exact steps that need to be taken?)
- What is the expected value(s)? How far is the data collected in the experiment from the expected value(s)?

Being able to detect these types of errors will depend on:

- A deep understanding of the related theoretical concepts.
- Understanding the common methodology used for the type of experiment being conducted.

The best places to find these errors will be in the methods and results section of the research paper.

Precision and accuracy 2.3.2.3

OVERVIEW

Precision and accuracy affect experimental results in different ways.

THEORY DETAILS

Although these terms both relate to the collection of data, they give us slightly different information. While **precision** describes the closeness of the measured values in an experiment, **accuracy** describes how close the measured values are to the true value. As seen in figure 4, measurements that are precise aren't necessarily accurate; in fact the measurements can be far away from the true value but still considered precise.

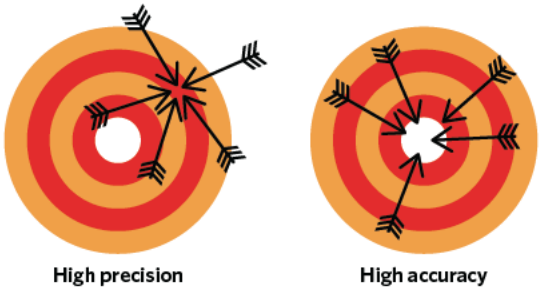


Figure 4 Comparison between precision and accuracy.

It is also possible to collect data that is highly accurate yet not precise if the values are not close to each other. We can therefore categorise data based on both accuracy and precision. This concept is illustrated in figure 5.

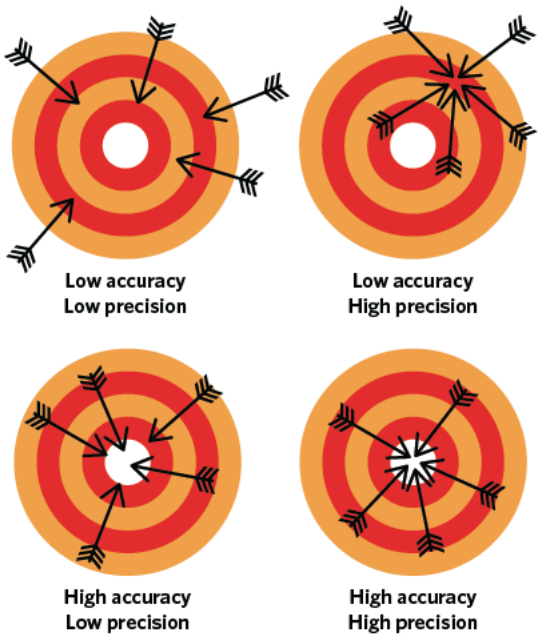


Figure 5 Illustration of the categorisation of data.

The accuracy of data collected during an experiment is affected by the presence of systematic errors in the experimental design, some of which may include:

- The purity of the substances used.
- Apparatus used is not cleaned properly.
- Measuring instruments are not calibrated prior to use.
- The experimental method is unclear.

To increase the accuracy of data collected, we need to try to remove any systematic errors that may be present in the experiment. To test the accuracy of our data, we can determine the measurement error of the results by finding the difference between the measured values and the true value.

Precision is more impacted by random errors. For experiments that require estimation of results, the measurements taken may not be very precise. This can occur during experiments where we are asked to identify changes in colour that are similar, or when trying to determine small changes in volume using large measuring cylinders.

The collection of data that is both accurate and precise is important as it can affect the conclusions made from an experiment. It also affects the **validity** of an experiment; the ability of the results obtained through the particular experimental design to address the aim and hypothesis of a research study. To make sure that results are valid, we need to make sure that:

- The appropriate measuring instruments are used to collect accurate results.
- Controlled variables are kept constant.
- The experimental methodology is actually able to test the hypothesis.
- Data is collected under controlled and appropriate conditions.

Significant figures in reports 2.3.9.4

OVERVIEW

Data is collected and expressed to various significant figures.

THEORY DETAILS

Significant figures represent the precision of a measurement tool. A measuring apparatus that expresses values in three significant figures is less precise than a tool that expresses values in four significant figures.

The general rules for counting significant figures are:

- Significant figures are counted starting from the first non-zero number (from left to right).
- Zeros that occur after we have started counting are only significant if they are followed by a non-zero number or a decimal place.

This concept can be shown in figure 6.

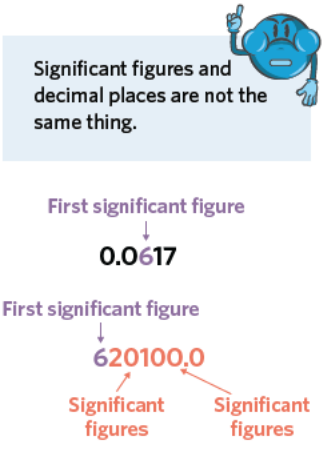


Figure 6 Consideration of significant figures.

Worked example 1

How many significant figures are found in the number 0.1104?

What information is presented in the question?

The number, 0.1104.

What is the question asking us to do?

Determine the number of significant figures.

What strategy(ies) do we need in order to answer the question?

1. Apply the rules to count the number of significant figures.

Answer

According to the rules, we start counting from the first non-zero number, which is the 1 immediately after the decimal place.

Since zeros that come after the first significant figure, that are followed by a non-zero number, are considered significant, the 0 after the 1 is considered significant.

Therefore, the number 0.1104 contains four significant figures.

When required to perform calculations, we need to express our answers to an appropriate number of significant figures. This depends on both the number of significant figures presented in the question, as well as the type of calculations that need to be performed:

- For addition/subtraction, answers are expressed with as many digits to the right of the decimal place according to the number with the least number of decimal places.
- For multiplication/division, answers are expressed with as many significant figures as the number with the least number of significant figures.

This is summarised in figure 7.

4.88 + 3.9 = 8.8
Answer is expressed to 1 decimal place (based on 3.9)

2.7 × 1.11 = 3.0
Answer is expressed to 2 significant figures (based on 2.7)

Figure 7 The strategies involved in expressing quantitative values from a calculation.

Worked example 2

As part of her calculations, Judee was to perform the following step:

$$= \frac{2.11}{1.3}$$

How many significant figures would her answer be expressed in?

What information is presented in the question?

The calculation to be performed.

What is the question asking us to do?

Determine the number of significant figures the answer should be expressed to.

What strategy(ies) do we need in order to answer the question?

1. Determine the type of calculation being performed.
2. Count the number of significant figures based on the appropriate rules.

Answer

This is a division calculation, therefore the answer will be expressed to the lowest number of significant figures given. For this question, the reference number will be 1.3, which contains two significant figures.

Therefore, the answer will be expressed to two significant figures.

Scientific notation

Sometimes numbers may be very large or very small which means that it is much easier to express them using scientific notation written as a decimal number multiplied by a power to the base of 10. The easiest way to do this is to:

- Move the decimal place in the number so that there is only one number to the left of the decimal point.
- Identify the number of significant figures required, and round the decimal number accordingly.
- Identify the number of times the decimal place has moved – this will be your exponent/power.



Sometimes the molar mass in the data book is the limiting factor with respect to significant figures.

Worked example 3

The diameter of the earth is approximately 12 742 km. Express this in scientific notation to three significant figures.

What information is presented in the question?

The diameter of the earth.

What is the question asking us to do?

Express the diameter of the earth to three significant figures using scientific notation.

What strategy(ies) do we need in order to answer the question?

1. Move the decimal place until there is only one number to the left of the decimal place.
2. Identify the number of significant figures required.
3. Write the answer in scientific notation.

Answer

12 742 can also be written as 12742.0. By moving our decimal, we can have a new number of:

$$1.2\,742.0$$

$$= 1.27420$$

Considering we want the answer to three significant figures, our third significant number will be seven. Since the fourth number is four, according to the rounding rules, our third significant figure will stay at seven.

Third significant figure → 1.27420 ← The number after 7 is less than 5, therefore our rounded number will be 1.27

$$= 1.27$$

Since we moved our decimal place four places to the left, our exponent/power will be four. Therefore the diameter can be expressed as:

$$1.27 \times 10^{4}\ \text{km}$$

Number of times the decimal point was moved

Detecting imprecise and inaccurate data in research

When examining other people's research, we can check for precision and accuracy by comparing the results obtained with other pieces of research, but also by comparing the results to what is expected. These expected results will be detected by studying the data obtained and represented in the results section of research.

Validity and the three R's 1.3.4.1 & 2.3.2.4

OVERVIEW

Errors can affect the ability of an experiment to answer the research question.

THEORY DETAILS

Unfortunately, not all experiments are considered valid. To ensure that experiments are able to test the aim and hypothesis, we need to make sure:

- We know what the independent variable is, and whether or not it relates to the question.
- We know the dependent variable, and whether or not it is what we need to measure to answer the question.
- We control all other variables.

There are three main ideas that are affected by the presence of errors.

Reproducibility

When experimental data is reproducible, the results collected by different experimenters using the same methodology will be within close proximity. Although these experiments will be conducted under slightly different conditions using different apparatus, the **reproducibility** can be increased when precise equipment is used.

Repeatability

Repeatability of experiments depends on whether or not relatively close values can be obtained when the experiment is repeated under the same conditions; that is, by the same experimenter, using the same apparatus, on the same day. The repeatability of experiments can be improved by increasing the number of measured values.

Reliability

Reliability of experiments depends on the consistency of results in the same experiment being conducted under varying conditions. Reliability can be increased by repeating experiments and averaging the results. Sometimes however, results can be outside the expected range. These are referred to as **outliers** and can often be investigated.

Making sure that data is valid and reliable ensures that the evidence collected is strong. Having strong evidence means that the findings from the research study is more robust and believable. In contrast, weak evidence that is laden with errors is less likely to be accepted by the scientific community.

Limitations of data 2.3.4.2

OVERVIEW

There are factors that can limit the ability to conduct experiments.

THEORY DETAILS

It is difficult to control every variable in an experiment. This limits the validity, accuracy and reliability of experimental results, making it sometimes difficult to identify the relationship between the independent and dependent variable. These limitations can be as a result of the design and methodology of the experiment, or even from the surrounding environment. Access to relevant apparatus and reagents as well as the depth of understanding of the related field can act as limitations to the research.

With respect to what we learn in VCE Chemistry, most of the data in the data book was developed from experiments that were conducted under particular conditions. Therefore, the information presented is true to that environment. However, when conducting our own experiments in the classroom, the conditions are not the same and therefore the data we use may not be in 100% accordance with what is stated in the data book.

Theory summary

- There are many different types of errors that can affect the quality of scientific research.
- Each error has a different effect on the overall research.
- Accuracy and precision measure different aspects of quantitative data.
- Significant figures allude to the precision of experimental data.
- The number of significant figures used to express answers depends on the mathematical operation being performed.
- Repeatability, reliability and reproducibility all refer to the ability of the results to be obtained by repeating the experiment under various conditions.
- There will always be variables that limit research investigations.

1D QUESTIONS

Theory review questions

Question 1

Which of the following is not a systematic error?

A Spilling a solution.

B Using a 100 mL measuring cylinder to measure all solutions of less than 10 mL.

Question 2

Which of the following is a random error?

A Forgetting to add a reagent.

B A parallax error.

Use the following diagram for questions 3–5.



Question 3

The diagram shown represents

A systematic errors.

B random errors.

Question 4

The diagram shown in question 3 shows data that is

A not precise.

B very precise.

Question 5

For the data shown to be accurate, the data points need to be

A as close as possible to the centre of the target.

B clustered together anywhere on the target.

Question 6

The number 150050.0 contains

A seven significant figures.

B five significant figures.

Use the following information for questions 7 and 8.

The two following calculations were performed:

1 $7.113 + 0.64$ **2** 0.51×1.996

Question 7

Which of the following statements is true?

A In calculation (1), the answer would be expressed to two decimal places.

B In calculation (1), the answer would be expressed to two significant figures.

Question 8

Which of the following statements is true?

A In calculation (2), the answer would be expressed to four significant figures.

B In calculation (2), the answer would be expressed to two significant figures.

Question 9

Experiments with reproducible results

A can be replicated by other experimenters to yield similar results.

B can be replicated by the same experimenter to yield similar results.

Question 10

Experiments are valid if

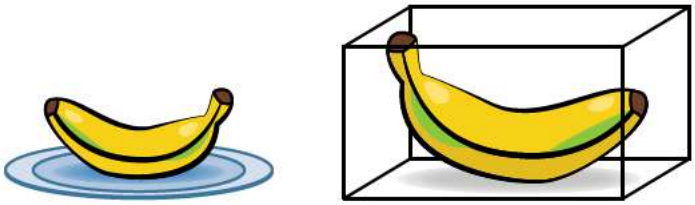
A all variables outside the independent (and dependent) variable are controlled.

B all variables are slightly adjusted throughout the experiment.

Deconstructed questions

Use the following information to answer questions 11–13.

Farquaad developed an experiment to identify the rate at which bananas ripen in an open area compared to in a closed area. To do so, he set up the following experiment:



Farquaad measured the time it took, in hours, from when the banana first started to turn brown. He checked the results every few hours during the day, but did not check it during the night.

Question 11 (1 MARK)

Which of the following steps will **not** improve the reliability of the experiment?

A Using bananas from the same bunch.

B Using the same type of bananas.

C Taking measurements at timed intervals.

D Changing the experimenter.

Question 12 (1 MARK)

Which of the following variables need to be kept constant?

A The size of the bananas.

B The temperature of the experiment.

C The timer used to measure the time.

D All of the above.

Question 13 (3 MARKS)

Based on the results of the experiment, Farquaad saw that the banana in the box began to ripen before the banana on the plate. Therefore, he concluded that bananas kept in a closed area ripen quicker than those left in an open environment. Comment on the validity of this conclusion.

Exam-style questions

Within lesson

Question 14 (4 MARKS)

For each of the descriptions, identify the type of error that has occurred.

Description	Type of error
Experimenter uses an uncalibrated thermometer to measure the heat of a reaction.	
Some reagent falls off a measuring glass as a student walks back to the bench.	
Student reads a measuring cylinder from the top of the meniscus.	
Student uses the wrong reagent in a single reaction.	

Question 15 (5 MARKS)

Ammonia is a strong cleaning agent found in many cleaning products. Depending on the purpose of the cleaner, different cleaners have different concentrations of ammonia. As part of an investigation to determine the concentration of ammonia, ammonia was reacted with different volumes of hydrochloric acid. It was expected that for every 20.00 mL of ammonia used there would be 15.00 mL of hydrochloric acid required.

The results from the test are provided.

Test number	Volume of HCl reacted (mL)
1	18.81
2	19.00
3	14.76
4	18.74

a Which test results are most precise? (1 MARK)

b Are the results accurate? Explain. (3 MARKS)

c The student wanted to find out the total volume of HCl used in the experiment. How many decimal places should the answer be expressed to? (1 MARK)

Question 16 (7 MARKS)

For their research investigation, a student wanted to conduct an experiment using a substance that is not commonly available in schools. As a consequence, they had to source a small sample from a nearby university.

a Identify two limitations of the experiment. (2 MARKS)

b The student followed methodology presented by a research paper studying a similar question in the same field. Despite this, the results obtained in the experiment and the results reported by the researchers in the article were very different. Comment on the reproducibility of the results. (2 MARKS)

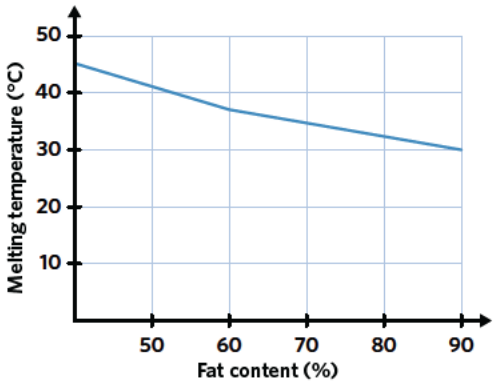
c Suggest a way in which the reproducibility can be improved. (1 MARK)

d What is the main difference between reproducibility and repeatability? (2 MARKS)

Multiple lessons

Question 17 (4 MARKS)

Balnea conducted an experiment to see how the fat content in chocolate affects the temperature at which the chocolate will melt.



a Identify the independent variable. (1 MARK)

b Identify an appropriate control group for the experiment. (1 MARK)

c The results obtained were plotted on a graph as shown. What conclusion can be drawn from the results? (1 MARK)

d Balnea wanted to calculate the amount of sugar present in the chocolate. To do so, 4.00 g samples of chocolate were required. How many significant figures would the answer be expressed in? (1 MARK)

Question 18 (4 MARKS)

Type 1 diabetes affects the body's ability to bring glucose into cells as there is not enough insulin. Insulin is responsible for helping to move glucose from the blood stream to cells. This means that most of the sugar isn't delivered to cells and stays in the blood. Therefore, those with type 1 diabetes need to closely monitor their blood glucose levels.

The different sugar content in foods can affect blood glucose in different ways. Thomas wanted to test different foods and their effect on his blood glucose. As part of this test, he ate a variety of different snacks with different glucose contents and measured his blood glucose every hour for four hours.

a Identify the dependent variable. (1 MARK)

b Write an appropriate hypothesis for this research. (1 MARK)

c At the end of the test, Thomas decided to calibrate his glucose meter and realised that his meter was measuring values 0.20 mmol/L less than the true value. Identify how this would affect his results. (2 MARKS)

Exam-style question hints

14: Different actions lead to different errors based on their effect on the data. **15a:** Precision describes how close values are to each other. **15b:** Accuracy describes how close values are to the true value. **15c:** The decimal places expressed depends on the data used. **16a:** Limitations can affect the ability of an experiment to be performed. **16b:** Reproducibility is measured in certain conditions. **16c:** improvements depend on the nature of the variable that needs to be optimised. **16d:** Reproducibility and repeatability differ in their experimental conditions. **17a:** The independent variable is deliberately manipulated. **17b:** Graphs show the relationship between the independent and dependent variable. **17c:** Expression of significant figures depends on the data used. **17d:** The number of significant figures depends on the question and the calculation performed. **18a:** The dependent variable is affected by the independent variable. **18b:** The hypothesis predicts how the independent variable will influence the outcome of the dependent variable. **18c:** Different errors affect results in different ways.

1E WRITING SCIENTIFIC MATERIAL

In this lesson, we will learn about how to share scientific information in a way that is easy to understand and that follows the conventions of scientific writing.

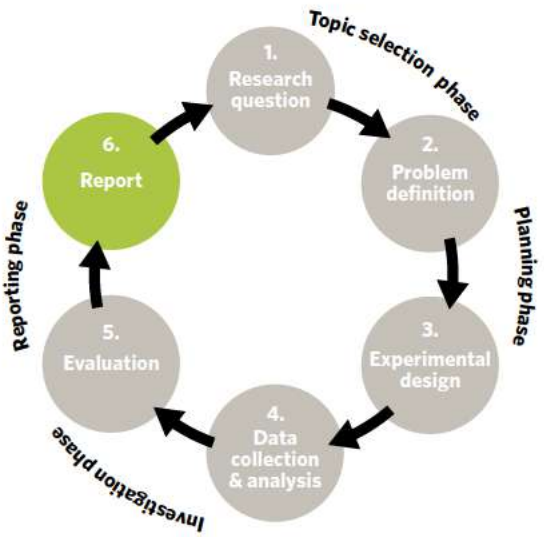
1A Scientific research | 1B Conducting an experiment | 1C Interpreting data | 1D Experimental factors affecting data | 1E Writing scientific material

Study design dot points

- the characteristics of effective science communication: accuracy of chemical information; clarity of explanation of chemical concepts, ideas and models; contextual clarity with reference to importance and implications of findings; conciseness and coherence
- the chemical concepts specific to the investigation: definitions of key terms; use of appropriate chemical terminology, conventions, units and representations
- the conventions of scientific report writing including chemical terminology and representations, symbols, chemical equations, formulas, units of measurement, significant figures and standard abbreviations
- the use of data representations, models and theories in organising and explaining observed phenomena and chemical concepts, and their limitations

Key knowledge units

Effective scientific communication	1.3.1.1
Chemical terminology, conventions and abbreviations	1.3.2.1 & 2.3.9.2
Units of measurement in reports	2.3.9.3
Representing ideas with models	1.3.1.2
Limitations of representations	1.3.3.2



Lesson links

This lesson builds on:

- 1C Interpreting data

It is important to represent data and ideas in an appropriate manner.

Effective scientific communication 1.3.1.1

OVERVIEW

Scientific communication needs to be concise.

THEORY DETAILS

The purpose of scientific articles and reports is to be able to convey a finding/message to the reader. Therefore, it is important to be concise and clear in the way we represent information without being verbose. Getting to the point is more important than sounding sophisticated. Some key considerations for effective communication in science include:

- Use short sentences that get to the point.
- Consistency of voice.
- Maintain a neutral stance and tone by avoiding subjective language.
- Use of visuals to support writing where necessary.
- Referencing of related theory.

Another important factor to consider when communicating scientific ideas, especially in a research article, is the appropriate formatting of the main text as well as any references used. For example, scientific research is written in a particular order with predetermined sections as shown in figure 1.

When we put this into the context of the water bottle experiment, we can develop a scientific poster to share our experiment as seen in figure 2.

The different sections of the poster convey different types of information, and together they provide a deeper understanding of an idea. Based on figure 2, we can see that a large portion of the poster is the discussion section. In fact, this is quite common in any scientific writing as this is the section where we bring all of the data and ideas together. The discussion section gives us an opportunity to elaborate on and explain the results, and to show how the research and its results contributes to the pool of knowledge relating to the same or similar ideas. Regardless of whether or not the results support the hypothesis, the analysis of the results and what they mean is the main purpose of the discussion. From this analysis, we are able to identify any future implications of the experiment.



Figure 1 Outline of a scientific write up.

Understanding that scientific ideas are built upon previous research, it's important to acknowledge the source of the information that formed the basis of the research. Depending on where scientific work is submitted to, the referencing requirements may vary. For example, some publications require references to be done in Harvard style, whereas others may require APA style. Information about the exact formatting required is available online.

Chemical terminology, conventions and abbreviations 1.3.2.1 & 2.3.9.2

OVERVIEW

Chemistry has its own language when it comes to representing information.

THEORY DETAILS

We know that chemistry has a particular language when presenting ideas. This includes the terms used, structural representations of chemicals and symbolic representations of elements. The most recognisable example is the use of symbols to represent the names of the chemical elements, for example the use of the symbol O to represent the element oxygen. Other examples include:

- The use of subscripts and superscripts to represent multiples or charges of atoms e.g. NH_3 and Mg^{2+}.
- The use of brackets in chemical formulas and structural formulas.
- The use of variables/symbols to represent chemical concepts e.g. N_A, n, M_r.
- Key terms representing multi-faceted chemical concepts e.g. compound and molecule.

Although there are certain terms and conventions that are used in different contexts, scientific concepts require specific conventions and terminology. Therefore, it is important that we are able to select these according to the context. At times, scientific research will include an appendix that outlines the conventions and abbreviations relevant to the research that is not commonly used elsewhere.

It is also important that anything included in the research is labelled appropriately. For example, tables and figures included in scientific writing all need to be accompanied by a table or figure legend. Furthermore, graphs need to be appropriately titled, with axis headings included with their appropriate units.

We will learn about these and other different representations throughout the lessons in this book.

Is it dangerous to drink water from a plastic water bottle that has been put in a hot environment?

Introduction

Bisphenol A (BPA), as shown in figure 1, is a monomer, a small building block, used to produce a strong plastic polymer known as polycarbonate (Vogel, 2009). Originally developed for medical uses, BPA is now used to produce strong plastics used in everyday products such as bottles, food containers and pipes. Due to its widespread use, there is growing concern over the health effects of using BPA in household plastic items.

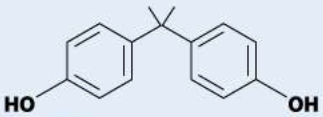


Figure 1 Structure of a bisphenol A (BPA) monomer.

In water-containing products such as drink bottles, BPA is known to leach into the water over time (Le et al, 2008). When consumed, this can result in the disruption of hormone-driven processes in the body (Diamanti-Kandarakis et al. 2009). This is potentially dangerous as there are many reactions in the body that are driven by hormones.

In order to be released into the water, BPA first needs to be hydrolysed. Due to this, factors such as temperature, pH conditions and microwaves can all affect the extent to which BPA is released into its surrounding environment.

This experiment aims to test the effect of temperature on the release of BPA into water. It's expected that if the breakdown of BPA from plastic depends on the temperature of the surrounding environment, then an increase in temperature would result in an increase in the concentration of BPA found in the water.

Materials

- 3 × 600 mL unopened BPA-containing bottles of water (from the brand Rolo Water)
- 3 × 600 mL unopened bottles of water in a BPA-free plastic bottle (from the brand Rolo Water)
- 1 × HPLC machine
- 0.50 ng/mL, 1.0 ng/mL, 1.5 ng/mL and 2.0 ng/mL BPA standards
- 1 × Micropipette
- 3 × Pipette tips
- 1 × Fridge
- 1 × Incubator



Method

Part A

Identification of the retention time of BPA.

1. Inject a 5 µL sample of a 2.0 ng/mL BPA standard into the HPLC machine.
2. Record the retention time of the most prevalent peak.
3. Repeat steps 1 and 2 with a fresh sample of the same standard.

Part B

Development of the BPA calibration curve.

1. Inject a 5 µL sample of 0.50 ng/mL BPA standard through the HPLC machine.
2. Record the area under the peak for the sample.
3. Repeat steps 1–2 for the 1.0 ng/mL, 1.5 ng/mL and 2.0 ng/mL standards.
4. Construct a calibration curve using the data collected.

Part C

Determination of the concentration of BPA.

1. Label each BPA-containing bottle A, B and C, and each BPA-free bottle D, E and F.
2. Place bottles A and D in a fridge at 4°C.
3. Place bottles B and E in an incubator set at 20°C.
4. Place bottles C and F in an incubator set at 30°C.
5. Leave the samples for 2 hours.
6. After the 2 hour period, inject a 5 µL sample of water from bottle A into the HPLC machine using the micropipette.
7. Record the area under the peak for sample A at the retention time determined for BPA in Part A.
8. Repeat steps 6 and 7 for bottles B–F.

Results

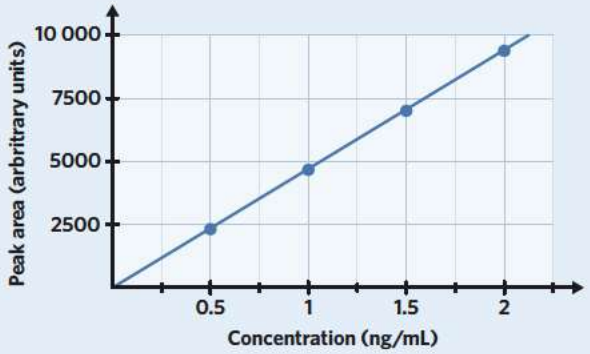


Figure 2.1 Calibration curve of known standards of BPA.

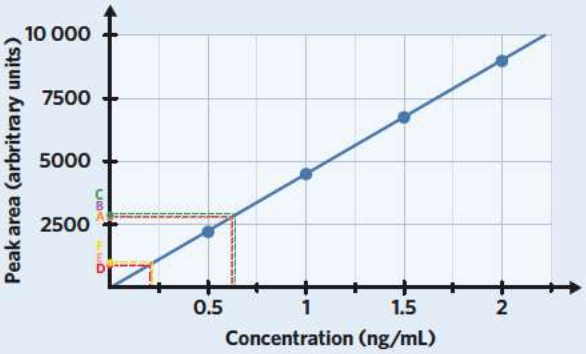


Figure 2.2 Determination of the concentration of BPA in samples A–F.

Figure 2 Scientific poster representing research on the effect of temperature on BPA concentration in water.

Results

Table 1 Concentration of BPA in each sample.

Sample	Peak Area (arbitrary units)	Concentration (ng/mL)
A BPA bottle at 4°C	2610	0.58
B BPA bottle at 20°C	2720	0.61
C BPA bottle at 30°C	2830	0.63
D BPA-free bottle at 4°C	101	0.022
E BPA-free bottle at 20°C	100	0.022
F BPA-free bottle at 30°C	101	0.022

Discussion

Compared to control samples D–F in which water was placed in a BPA-free bottle, the presence of BPA in the water bottles resulted in a greater concentration of BPA in the water samples. For samples A–C, it can be seen that an increase in temperature is correlated with an increase in the concentration of BPA measured. The samples placed in the 30°C and 20°C environments and in the fridge measured BPA concentrations in ng/mL of 0.63, 0.61 and 0.58 respectively.

BPA can be released into the contents of water bottles when the plastic begins to degrade. As shown in previous studies, an increase in temperature results in an increase in the degradation of plastic (Biedermann-Brem & Grob., 2008). In particular, an increase in temperature results in the increase in the rate at which the plastics are degraded. In the context of BPA-containing plastic water bottles, this could lead to an increase in the rate at which BPA is released into the water. This trend is shown in the results of this experiment. As can be seen in the experimental results, there was a greater concentration of BPA found in the sample subjected to the highest temperature (sample C) given that all samples were tested over the same period of time. It is possible then, that at higher temperatures, there could be a greater release of BPA from plastic, resulting in a higher concentration of BPA recorded.

Future implications

As BPA is found in a number of different plastic materials, the results of this experiment can help consumers make more informed decisions about the way they currently store goods.

Limitations

As there are so many different types of plastics available, the results of the study are specific to only Rolo Water.

Conclusion

In the same amount of time, an increase in temperature results in a slight increase in the concentration of BPA found in water of BPA-containing water bottles.



Refrences

Biedermann-Brem and K. Grob (2008). Release of bisphenol A from polycarbonate baby

bottles; water hardness as the most relevant factor. Eur. Food res. Technol., 228(5):679–684

Diamanti-Kandarakis, E., Bourguignon, J. P., Giudice, L. C., Hauser, R., Prins, G. S., Soto, A. M., Zoeller, R. T., & Gore, A. C. (2009). Endocrine-disrupting chemicals: an Endocrine Society scientific statement. Endocrine reviews, 30(4), 293–342.

Le, H. H., Carlson, E. M., Chua, J. P., & Belcher, S. M. (2008). Bisphenol A is released from polycarbonate drinking bottles and mimics the neurotoxic actions of estrogen in developing cerebellar neurons. Toxicology letters, 176(2), 149–156.

Vogel S. A. (2009). The politics of plastics: the making and unmaking of bisphenol a 'safety'. American journal of public health, 99 Suppl 3(Suppl 3), S559–S566.

Units of measurement in reports 2.3.9.3

OVERVIEW

Different units are used to represent different measurements.

THEORY DETAILS

In chemistry, we work with various different types of data ranging from amount, mass, volume and temperature, all of which are measured in different units. Figure 3 shows some common units that can be found in the VCAA data book.

Measured value	Conversion
0°C	273 K
100 kPa	750 mm Hg or 0.987 atm
1 litre (L)	1 dm^3 or 1×10^{-3} m^3 or 1×10^3 cm^3 or 1×10^3 mL

Metric (including SI) prefixes	Scientific notation	Multiplying factor
giga (G)	10^9	1 000 000 000
mega (M)	10^6	1 000 000
kilo (k)	10^3	1000
deci (d)	10^{-1}	0.1
centi (c)	10^{-2}	0.01
milli (m)	10^{-3}	0.001
micro (μ)	10^{-6}	0.000001
nano (n)	10^{-9}	0.000000001
pico (p)	10^{-12}	0.000000000001

Figure 3 Common units used in chemistry as outlined by the VCAA data book.

This is not a complete list of all of the units that we will use. The units used will depend on the type of information that we are collecting and trying to convey, and therefore, they vary contextually. We will come across more units as we work through this course.

Representing ideas with models 1.3.1.2

OVERVIEW

Chemical concepts can also be represented visually.

THEORY DETAILS

Models are often used to represent chemical concepts in a way that helps to make it easier to understand conceptual theory. Visual models are the most commonly used and perhaps most useful when conveying conceptual information. For example, the ball-and-stick model in figure 4 is a common representation of molecules and their bonds, and is often used in resources to demonstrate the complexities of chemical bonds.

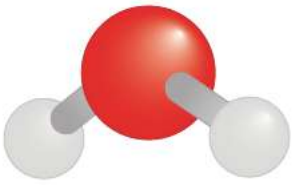


Figure 4 A ball-and-stick model representation of a water molecule.

Other visual representations include flow charts and diagrams such as the example in figure 5, which functions to summarise the overall steps involved in a process.

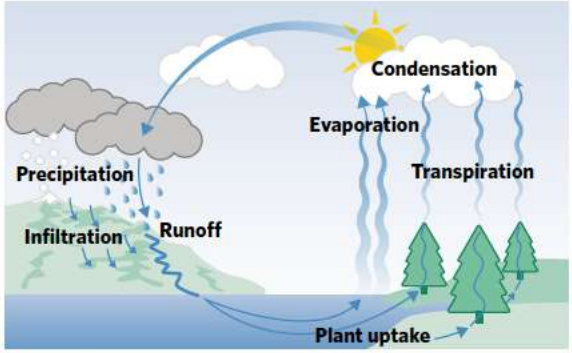


Figure 5 The water cycle.

Other commonly used visual representations include graphs and tables.

Models do not just have to be visual, but can also be worded, or even formulaic. Chemical equations and formulas are also examples of models used in chemistry as they are different ways of representing concepts.

Limitations of representations 1.3.3.2

OVERVIEW

Representations often do not show the complete picture.

THEORY DETAILS

Models and figures can make it easy to understand the general idea of chemical concepts, however, they are sometimes unable to illustrate the idea in detail. For example, with the water cycle shown in figure 5, we can identify the major components in the water cycle. Consequently, we cannot see what actually happens at each stage. The addition of the smaller details would make the model too complicated and hard to understand. Therefore, we need to understand that models are often used to summarise or demonstrate a particular idea, not to show all of the details involved in a concept.

These concepts, along with all of the other concepts covered in the chapter, allow us to both develop our own research as well as critique other experiments during our research journey. The outcome of this can be shown in the poster related to our water bottle experiment, as shown in figure 2.

Theory summary

- Being clear and concise is important when communicating scientific ideas.
- Different concepts contain different terminology and conventions.
- The units of measurement used depends on the type of information that is being collected.
- Chemical concepts can be represented visually, however, these can be limited in their ability to show the idea in great detail.

1E QUESTIONS

Theory review questions

Question 1

Scientific writing is supposed to be

A verbose and sophisticated.

B simple and clear.

Question 2

The following are examples of how water cycles through the environment. Which of the two statements would most likely be preferred when summarising this concept?

A Water exists in many forms and can therefore be seen to move through different stages in the environment.

B Due to the structural composition of water, which impacts its physical and chemical properties, it is able to interact with the environment in such a way that renders it capable of moving through different environments in different forms.

Question 3

A scientific write up must include

A an acknowledgment section.

B a methods section.

Question 4

Chemical terminology is used to describe

A a multi-process concept only.

B both a multi-process concept and single ideas.

Question 5

Units

A can be used to represent multiple different concepts.

B are often specific to what it is that is being measured.

Question 6

The most suitable unit to measure distance would be

A minutes.

B metres.

Question 7

A chemical equation can be considered as a model.

A True

B False

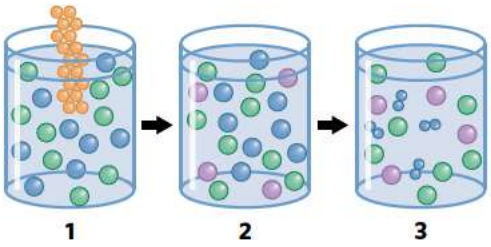
Question 8

Which of the following is an example of a visual model.

A



B



Question 9

Visual models cannot

A include words.

B show all details related to a concept.

Question 10

The image provided shows the different ways in which heat can be transferred. From this, we can see that

A convection involves the transfer of heat through the movement of particles of different kinetic energy.

B there are three main types of heat transfer.

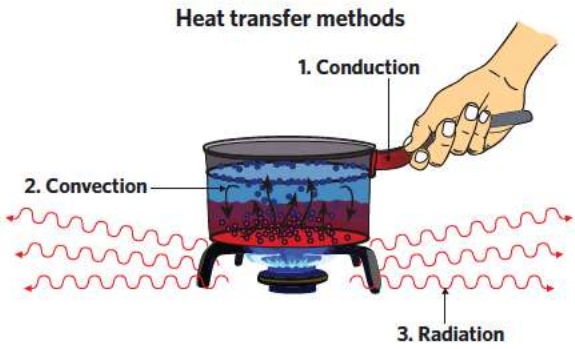


Image: udaix/Shutterstock.com

Deconstructed questions

Use the following information to answer questions 11–13.

The following are excerpts from an article titled 'What is a virus? How do they spread?'

> **Are viruses alive?**
>
> Viruses rely on the cells of other organisms to survive and reproduce, because they can't capture or store energy themselves. In other words they cannot function outside a host organism, which is why they are often regarded as non-living.
>
> Outside a cell, a virus wraps itself up into an independent particle called a virion. The virion can 'survive' in the environment for a certain period of time, which means it remains structurally intact and is capable of infecting a suitable organism if one comes into contact.
>
> **How do viruses spread?**
>
> Once a person is infected with a virus, their body becomes a reservoir of virus particles which can be released in bodily fluids – such as by coughing and sneezing – or by shedding skin or in some cases even touching surfaces.
>
> The virus particles may then either end up on a new potential host or an inanimate object. These contaminated objects are known as fomites, and can play an important role in the spread of disease.
>
> *Source: theconversation.com/what-is-a-virus-how-do-they-spread-how-do-they-make-us-sick-133437*

Question 11 (1 MARK)

The style of scientific writing above can be described as being

A complicated.
B clear.
C difficult to understand.
D sophisticated.

Question 12 (1 MARK)

Which of the following terms used in the article could be considered as a science-specific term related to viruses?

A Non-living
B Reservoir
C Virion
D Contaminated

Question 13 (3 MARKS)

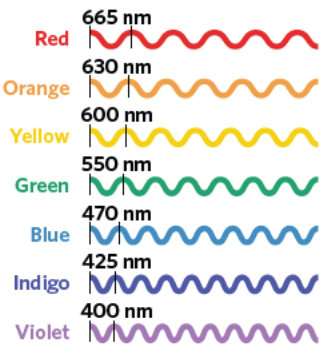
Construct a visual model to show the process of how a handshake can help spread viruses.

Exam-style questions

Within lesson

Question 14 (3 MARKS)

The light that we see is a combination of different colours.



a What is the main message conveyed by the visual model provided? (1 MARK)

b What is the unit of measurement used to measure the wavelength of light? (1 MARK)

c Identify the main limitation with this visual representation. (1 MARK)

Multiple lessons

Question 15 (3 MARKS)

Olaf wanted to conduct an experiment to see the effect of sunlight on the rate of growth of carrots. The results of the experiment are shown with the units omitted.

Test	Length of sun exposure	Length of carrot
A	0	0.15
B	1	15
C	2	24
D	3	29

a Identify the most appropriate unit for the data relating to the length of sun exposure. (1 MARK)

b Which of the tests would be considered as the control group? (1 MARK)

c Identify the dependent variable. (1 MARK)

Question 16 (3 MARKS)

Medusa received a crystal growing kit as a gift and wanted to figure out how to grow the largest crystals possible. The instructions for the kit are as follows:

1 Add boiling water to a crystal growing tray.

2 Stir mixture until everything is dissolved.

3 Place rocks found in the kit at the bottom of the tray.

The kit contained multiple different sized rocks that could be used to make crystals. The three crystal mixtures included in the kit contained were identical.

a Identify the independent variable. (1 MARK)

b Write an appropriate aim for the experiment. (1 MARK)

c Medusa wanted to write a poster to share the findings with other classmates. As part of this, outline the main sections that must be included in the poster. (1 MARK)

Exam-style question hints

14a: Different colours are represented differently. **14b:** Visual representations need to be labelled appropriately. **14c:** Visual representations often do not include the reasoning for certain concepts. **15a:** Units are appropriate for the data being measured. **15b:** Controls help to prove that the reason for change is due to the independent variable. **15c:** The dependent variable is affected by the independent variable. **16a:** The independent variable is the variable deliberately manipulated by the experimenter. **16b:** An aim outlines the purpose of the experiment. **16c:** Scientific writing has a predetermined format.

CHAPTER 1 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Use the following information to answer questions 1 and 2.

Dean conducted an experiment to determine people's emotional response to the use of face masks.

Question 1 (1 MARK)

The data collected would be considered as

A qualitative data.
B quantitative data.
C descriptive data.
D concrete data.

Question 2 (1 MARK)

The following conclusion was drawn from the results obtained:

'As the use of face masks increased, the happiness of people decreased.'

The relationship between the use of face masks and happiness can be described as being

A causal.
B correlated.
C unrelated.
D captivated.

Question 3 (1 MARK)

The purpose of a control is to

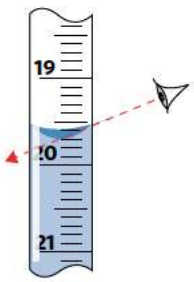
I improve the accuracy of results.
II improve the validity of results.
III show that any changes were due to the independent variable.

A I only
B III only
C II and III only
D II only

Question 4 (1 MARK)

The error illustrated by the diagram would have the greatest effect on

A the precision of results.
B the accuracy of results.
C the repeatability of an experiment.
D the conclusion of an experiment.



Question 5 (1 MARK)

Which of the following is **not** considered as a limitation?

A Access to appropriate reagents.
B Misinterpretation of the methodology.
C Access to information.
D The field of research is relatively new.

Question 6 (1 MARK)

According to the MSDS provided, the substance is dangerous as it is

A considered 3 on the richter scale.
B really heavy.
C able to freeze easily.
D can cause irritation.

Hazards	
MSDS	External MSDS
R-phrase	R11, R36, R66, R67
S-phrase	S16, S26, S33
Main hazards	Flammable (F) Irritant (Xi)
NFPA 704	

Question 7 (1 MARK)

Forgetting to perform a step outlined in the methodology of an experiment would be considered as

A a systematic error.
B an obstructive error.
C a random error.
D bias.

Question 8 (1 MARK)

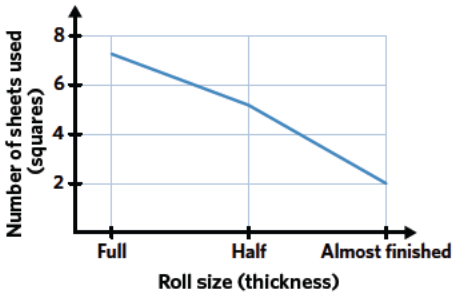
At a training session one day, Anthony measured the time it took him to run 400 m. The controlled variables in this experiment would include:

A the timer used, trial times and the shoes worn.
B the time recorded, the running route and the apparel worn.
C the distance, the running route and the rest period.
D the water intake, the rest period and his pace.

Question 9 (1 MARK)

From the graph provided, it can be seen that

A the number of sheets of paper used is constant.
B the number of sheets used increases as the size of roll decreases.
C the larger the roll, the less sheets used.
D the smaller the roll, the less sheets used.



Question 10 (1 MARK)

Kat was practicing her basketball shots and used green paint to map each of her three shots. Based on the results, her shooting can be described as being

A accurate but not precise.
B precise but not accurate.
C both precise and accurate.
D neither accurate nor precise.



Image: Ajay Shrivastava/Shutterstock.com

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (4 MARKS)

Oobleck is a substance that is unusual. As a liquid, it can be poured however when there is a force acting on it, it behaves like a solid. Ooblek is easy to make and consists of corn flour and water.

Shuka wanted to test the effect of the content of corn flour on the 'hardness' of the oobleck.

a Identify the independent variable. (1 MARK)

b What type of data is being collected in this experiment? (1 MARK)

c As part of the experiment, Shuka tested the hypothesis using the following samples:

Sample 1	Sample 2	Sample 3
1.0 g cornflour + 100 mL of water	2.0 g of cornflour + 200 mL of water	3.0 g of cornflour and 300 mL of water

Comment on the validity of the experiment. (2 MARKS)

Question 12 (6 MARKS)

To reduce the impact of diet-related diseases, researchers have tried to develop and suggest a taxation system that helps Australians to become healthier. The model that the researchers developed is shown.

Intervention	Tax or subsidy	Sources of assumptions
Saturated fat tax	$1.37/100 g of saturated fat	Tax on saturated fat content in foods with >2.3% saturated fat, excluding drinking milk
Excess salt tax	$0.30/100 g of sodium	Tax on sodium in excess of Australian maximum recommended levels, excluding fresh fruits, vegetables, meats and dairy products
Sugar-sweetened beverage tax	$0.47/l	Tax on sugar-sweetened soft drinks, energy drinks, cordials, and fruit drinks
Fruit and vegetable subsidy	$0.14/100 g	Subsidy on all fresh and preserved fruits and vegetables
Sugar tax	$0.94/100 ml of ice cream; $0.85/100 g of sugar	Tax on ice cream containing > 10 g of sugar per 100 g of ice cream; tax on sugar content in excess of 10 g per 100 g of all other products, excluding fresh fruits, vegetables and unflavoured dairy products

a Write an hypothesis for this experiment. (1 MARK)

b Suggest what effect would this tax system have on food consumption? (2 MARKS)

c What are the main concerns related to this system? (3 MARKS)

Question 13 (8 MARKS)

Jasper loves playing with bubbles. After running out of his bubble solution, he decided to develop a solution that would make the largest bubbles.

After conducting research, Jasper realised that bubble solutions are made from different types of substances and water. Therefore, he decided to test three different solutions containing dishwashing detergent, glycerin and corn syrup.



Image: David Carillet/Shutterstock.com

Method

1. Add 5mL of dishwashing detergent to 200 mL of water.
2. Stir solution.
3. Place the bubble wand into the solution for 3 seconds.
4. Pull the bubble wand out of the solution and wave the hand once to make bubbles.
5. Take a photo of the bubbles formed.
6. Measure the diameter of the largest bubble in the photo.
7. Repeat for glycerin and corn syrup.

a Identify an appropriate control for this experiment. (1 MARK)

b The results of the experiment are shown.

	Diameter of the largest bubble (cm)			
Substance	**Test 1**	**Test 2**	**Test 3**	**Average**
Detergent	5	4.8	3	4.3
Glycerin	7.1	7.2	7	7.1
Corn syrup	7.5	7.6	7.5	7.5

i Would a bar graph, column graph or line graph be most suitable to represent the results shown in the table? (1 MARK)

ii Which substance gave the most precise results? (1 MARK)

iii Comment on the reproducibility of the results. (3 MARKS)

iv Based on the data, what can be concluded about the experiment? (2 MARKS)

Question 14 (6 MARKS)

Scientist Émilie du Châtelet proposed a law stating that energy cannot be created or destroyed; it can only be transformed or transferred. This principle is still believed to be true to this day.



Image: Visual Generation/Shutterstock.com

a Consider the scenario shown, involving the use of a coal barbecue to cook food. Based on this scenario, a student made the following statement about the law of conservation of energy.

'Energy can be lost because the chemical energy that was available in the coal was used to start a fire to cook food. However, some of the heat escaped the barbecue and was not able to be used to cook the food.'

Comment on the accuracy of the statement made. (2 MARKS)

b To prove the theory, the student wanted to develop an experiment that could measure the energy in the coal, the barbecue, the food and the air to see if the total amount of energy would be the same.

i Identify a limitation of the experiment. (1 MARK)

ii List three variables that would need to be controlled. (3 MARKS)

Question 15 (6 MARKS)

The taste buds on our tongue are sensitive to different tastes. The amount of food required to 'taste' something can vary between people, but can also be trained. For example, we can train our taste buds to increase the threshold of sourness that we can handle before we register the sour taste by consuming foods with a high sourness rating.

In an experiment to test this hypothesis, researchers wanted to find out whether or not having subjects consume varying concentrations of vinegar prior to eating a piece of lemon would reduce a person's reaction to consuming the lemon.

a Identify the independent variable. (1 MARK)

b To measure the ability of a person to consume the lemon, the experiments measured both facial expressions and a self-rating score of one to five, where five corresponded to 'very sour'. The results of the experiment are shown.

Concentration of vinegar (M)	1.0	2.0	3.0	4.0
Facial expression				
Self rating	3	2	4	4

i Describe a key ethical concern in this experiment. (2 MARKS)

ii Comment on the validity of the experiment. (2 MARKS)

iii What conclusion can be drawn from the experiment? (1 MARK)

UNIT

1 How can the diversity of materials be explained?

The development and use of materials for specific purposes is an important human endeavour. In this unit students investigate the chemical properties of a range of materials from metals and salts to polymers and nanomaterials. Using their knowledge of elements and atomic structure students explore and explain the relationships between properties, structure and bonding forces within and between particles that vary in size from the visible, through nanoparticles, to molecules and atoms.

Students examine the modification of metals, assess the factors that affect the formation of ionic crystals and investigate a range of non-metallic substances from molecules to polymers and giant lattices and relate their structures to specific applications.

Students are introduced to quantitative concepts in chemistry including the mole concept. They apply their knowledge to determine the relative masses of elements and the composition of substances. Throughout the unit students use chemistry terminology including symbols, formulas, chemical nomenclature and equations to represent and explain observations and data from experiments, and to discuss chemical phenomena.

A research investigation is undertaken in Area of Study 3 related to one of ten options that draw upon and extend the content from Area of Study 1 and/or Area of Study 2.

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UNIT 1

AOS 1

How can knowledge of elements explain the properties of matter?

In this area of study students focus on the nature of chemical elements, their atomic structure and their place in the periodic table. They review how the model of the atom has changed over time and consider how spectral evidence led to the Bohr model and subsequently to the Schrödinger model. Students examine the periodic table as a unifying framework into which elements are placed based upon similarities in their electronic configurations. In this context students explore patterns and trends of, and relationships between, elements with reference to properties of the elements including their chemical reactivity.

Students investigate the nature of metals and their properties, including metallic nanomaterials. They investigate how a metal is extracted from its ore and how the properties of metals may be modified for a particular use. Students apply their knowledge of the electronic structures of metallic elements and non-metallic elements to examine ionic compounds.

They study how ionic compounds are formed, explore their crystalline structures and investigate how changing environmental conditions may change their properties.

Fundamental quantitative aspects of chemistry are introduced including the mole concept, relative atomic mass, percentage abundance and composition by mass and the empirical formula of an ionic compound.

Outcome 1

On completion of this unit the student should be able to relate the position of elements in the periodic table to their properties, investigate the structures and properties of metals and ionic compounds, and calculate mole quantities.

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UNIT 1 AOS 1, CHAPTER 2

02 Atoms, particles and molecules

2A The atom

2B Defining elements

2C The periodic table - part 1

2D The periodic table - part 2

2E Relative masses

2F The mole - part 1

2G The mole - part 2

2H Chemical formulas and equations

Key knowledge

- the relative and absolute sizes of particles that are visible and invisible to the unaided eye: small and giant molecules and lattices; atoms and subatomic particles; nanoparticles and nanostructures
- the definition of an element with reference to atomic number; mass number; isotopic forms of an element using appropriate notation
- spectral evidence for the Bohr model and for its refinement as the Schrödinger model; electronic configurations of elements 1 to 36 using the Schrödinger model of the atom, including s, p, d and f notations (with copper and chromium exceptions)
- the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic/non-metallic character and reactivity) of elements
- the relative isotopic masses of elements and their representation on the relative mass scale using the carbon-12 isotope as the standard; reason for the selection of carbon-12 as the standard
- determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)
- the mole concept; Avogadro constant; determination of the number of moles of atoms in a sample of known mass; calculation of the molar mass of ionic compounds
- experimental determination of the empirical formula of an ionic compound
- determination of empirical and molecular formulas of organic compounds from percentage composition by mass and molar mass
- the conventions of scientific report writing including chemical terminology and representations, symbols, chemical equations, formulas, units of measurement, significant figures and standard abbreviations

Image: vchal/Shutterstock.com

2A THE ATOM

In this lesson, we will be learning about the size of different objects and the various models of the atom.

2A The atom | 2B Defining elements | 2C The periodic table - part 1 | 2D The periodic table - part 2 | 2E Relative mass | 2F The mole - part 1 | 2G The mole - part 2 | 2H Chemical formulas and equations

Study design dot points

- the relative and absolute sizes of particles that are visible and invisible to the unaided eye: small and giant molecules and lattices; atoms and subatomic particles; nanoparticles and nanostructures
- spectral evidence for the Bohr model and for its refinement as the Schrödinger model; electronic configurations of elements 1 to 36 using the Schrödinger model of the atom, including s, p, d and f notations (with copper and chromium exceptions)

Key knowledge units

Size matters	1.1.1.1
Nanomaterials	1.1.1.2
The atom	1.1.1.3
The Bohr model	1.1.3.1
The Schrödinger model	1.1.3.2

Key terms and definitions

Scale range of values used for measuring size or other attributes

Nanoscale measurements on the 10^{-9} m scale

Nanoparticles particles on the 10^{-9} m scale

Adsorption attachment of a molecule to a surface

Atom smallest unit of matter

Subatomic particle particles that exist inside the atom

Electron light negatively charged particle that exists outside the nucleus of an atom

Proton light positively charged particle inside the nucleus of an atom

Neutron dense neutrally charged particle inside the nucleus of an atom

Nucleus region at the centre of an atom that contains protons and neutrons

Nucleons subatomic particles found in the nucleus

Lattice regular three-dimensional arrangement of atoms

Energy shells or energy level orbits, containing different levels of energy, around the nucleus of an atom where electrons are found according to the Bohr model

Emission spectra band of frequencies emitted by an atom due to electrons returning to the ground state from an excited state

Ground state lowest energy level that an electron occupies

Excited state temporary energy level that an electron occupies once it has absorbed energy

Electron configuration arrangement of electrons in shells and/or subshells

Quantum mechanics field of study which looks at the behaviour of very small objects

Orbitals regions with the highest probability of finding electrons

Pauli exclusion principle rule that states an orbital can't hold more than two electrons

Aufbau principle rule that states subshells are filled by electrons from the lowest to the highest energy level

Size matters 1.1.1.1

OVERVIEW

The atom is the basic building block of matter and is invisible to the naked eye.

THEORY DETAILS

One of the major developments in science is the ability to measure the size of objects. Scientists have created a **scale** which compares the sizes of different objects as shown in figure 1.

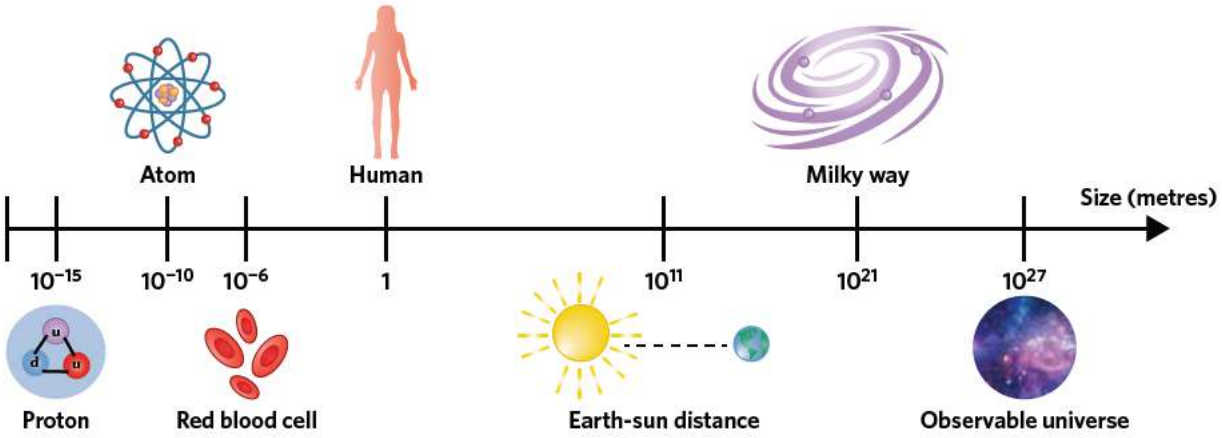


Image: van van/Shutterstock.com

Figure 1 Scale of objects.

Figure 1 shows how a red blood cell which is invisible to the naked eye is smaller than a tennis ball by many orders of magnitude. Chemistry relies on specific details regarding size and scale, and when this involves objects that cannot be seen with the naked eye, this measuring system becomes very important.

The measuring system that we use has a base unit of metres (m). This can then be adapted by using a metric prefix which is a letter in front of the m which changes the multiplying factor. The prefixes can be found in your data book for Chemistry Units 3 & 4 in table 1.

Table 1 Prefixes and their respective multiplying factor.

Metric (including SI) prefixes	Scientific notation	Multiplying factor
Giga (G)	10^{9}	1 000 000 000
Mega (M)	10^{6}	1 000 000
Kilo (k)	10^{3}	1000
Deci (d)	10^{-1}	0.1
Centi (c)	10^{-2}	0.01
Milli (m)	10^{-3}	0.001
Micro (μ)	10^{-6}	0.000001
Nano (n)	10^{-9}	0.000000001
pico (p)	10^{-12}	0.000000000001

Source: VCAA data book

These prefixes are important in chemistry because sometimes numbers are too big or too small to manage when using just the unit of metres. Therefore, prefixes are added to the units to make them easier to manage. Converting between units is an important skill because sometimes calculating the amount of substances present from an equation requires unit conversions.

Worked example 1

The average length of a blue whale is 29.9 metres. What is its length in nanometres?



What information is presented in the question?

The length in metres is 29.9.

What is the question asking us to do?

Convert metres to nanometres.

What strategy(ies) do we need in order to answer the question?

1. Find the multiplying factor that belongs to the nano prefix.
2. Use this multiplying factor to convert metres to nanometres.
3. Write the answer with correct scientific notation.

Answer

The nano prefix requires a multiplying factor of 10^{-9}

$1\ m = 1 \times 10^{9}\ nm$

$29.9\ m = 29.9 \times 10^{9}\ nm$

$= 2.99 \times 10^{10}\ nm$

The length of the whale is $2.99 \times 10^{9}\ nm$

Nanomaterials 1.1.1.2

OVERVIEW

Nanomaterials are objects that are invisible to the naked eye and are a relatively new discovery in chemistry.

THEORY DETAILS

In order to understand just how small these nanomaterials are, we need to use the **nanoscale**, also known as the nanoscopic scale. One nanometre (1 nm) is equal to 10^{-9} m, or 0.000000001 metres. This means that you could fit one billion nanometres into a metre, or a million into a centimetre. Alternatively, your smallest fingernail is approximately 0.5 mm thick, which still fits about 50000 nm in it.

Due to the size of nanomaterials, they can have significantly different properties than the material which they were made from or constructed with. One example that illustrates the diversity of nanomaterial properties is a carbon nanotube, depicted in figure 2. These, unlike what we normally think about carbon containing compounds like coal, are flexible and very strong.

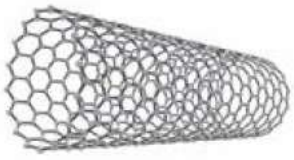


Image: Evannovostro/Shutterstock.com

Figure 2 Structure of a carbon nanotube.

Nanoparticles are a specific type of nanomaterial which are spherical in shape, and have various uses in medical, optical and electrical fields. These nanoparticles have a large surface area in comparison to their volume. This is important due to its effect on adsorption.

Adsorption is the process in which particles stick to the outside of an object. You may have seen adsorption happen when a drop of water sticks to a shower panel. Nanoparticles are very effective at adsorption because of their large surface area. Due to this, nanoparticles can adsorb a large amount of particles for various functions, including the transport or removal of unwanted substances, like cancer cells. In addition, some nanoparticles are used as catalysts, which speed up the rate of reactions.

These varying properties offer exciting new ways for scientists to solve problems. For example, research is being conducted into cancer-fighting nanoparticles as shown in figure 3. Here, the nanoparticles have adsorbed to a cancer cell.

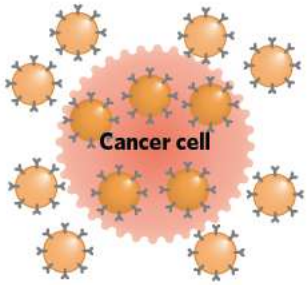


Figure 3 Cancer-fighting nanoparticles.

The atom 1.1.1.3

OVERVIEW

Atoms have a basic general structure.

THEORY DETAILS

Throughout the history of science, there have been many different scientific models which describe how matter is arranged and created. Ever since the ancient Greek philosopher Democritus described **atoms** as the smallest particles which are invisible, this model has changed and transformed. We will be learning about the different models of the atom later in this lesson. Figure 4 is a timeline of all the different models of the atom.

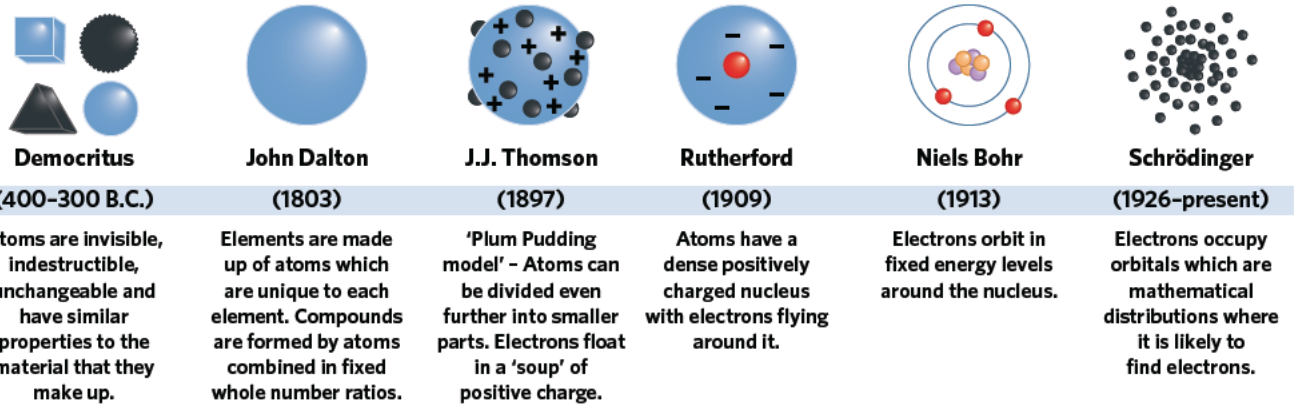


Figure 4 Timeline of different models of the atom.

It is important to note that the models shown in figure 4 are not definitive representations of reality. Over time, they have and will constantly adapt and evolve. For example, John Dalton's model of the atom stated that atoms are indivisible units of matter and elements consist of a single type of particle. While his theory is viewed as mostly correct today, scientists have now discovered that atoms can be further divided into **subatomic particles**. This highlights the importance of scientific research, and how the model of the atom is constantly evolving as new discoveries are made.



New subatomic particles like quarks and muons are constantly being discovered by physicists!

Subatomic particles come in three main forms: **protons, neutrons and electrons**. Each of these subatomic particles are found in different areas in the atom and vary in size. A basic visualisation of the atom can be seen in figure 5.

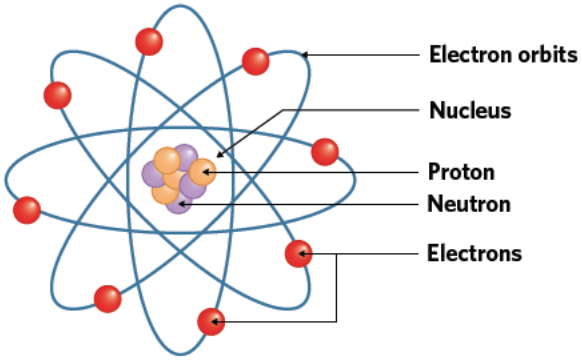


Figure 5 Atomic structure with protons, neutrons and electrons labelled.

As we can see, the protons and neutrons can be found at the centre of an atom, which make up the **nucleus** of an atom. Protons are positively charged particles, whereas neutrons are neutral and therefore do not carry any charge. These particles can also be referred to as **nucleons**. Surrounding the nucleus are the electrons; negatively charged particles that are constantly moving around the nucleus. It is the attraction between the positive protons in the nucleus and the negative charge of the electrons that holds the electrons to the atom.

In general, for neutral atoms, the number of electrons is equal to the number of protons to cancel out the charges of each subatomic particle. In terms of their mass, electrons are the lightest of the subatomic particles, whereas protons and neutrons are much heavier and possess similar masses. These properties are summarised in table 2.

Table 2 Properties of protons, neutrons and electrons.

Particle	Charge	Mass (kg)
Proton	+1	1.673×10^{-27}
Neutron	0	1.675×10^{-27}
Electron	−1	9.109×10^{-31}

In order to create all that we see (and sometimes don't see) in the world, single atoms come together to make bigger substances. Take a human body for example. As seen in figure 6, our most basic building block is the atom, which comes together to form molecules and progressively forms larger groups until a whole human being has been formed.

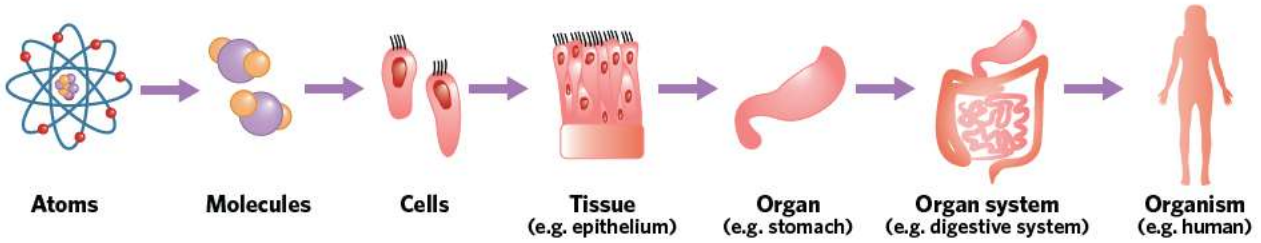


Figure 6 The building blocks of life.

Atoms can arrange themselves in different ways to create different structures, such as **lattices**, which are repeated three-dimensional arrangements of atoms that are bonded together. Some examples of lattices include graphite or graphene as shown in figure 7, which will be explored later in the course.

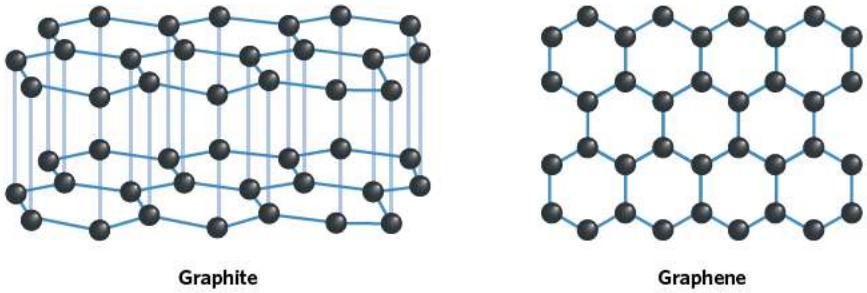


Figure 7 Visual representations of graphite and graphene lattices.

The Bohr model 1.1.3.1

OVERVIEW

The Bohr model of the atom states that electrons orbit around the nucleus in fixed paths.

THEORY DETAILS

The Bohr model shows that the atom consists of a dense, positively charged nucleus at the centre of the atom, with negatively charged, light electrons circling around the nucleus in fixed orbits as seen in figure 8. Bohr's theory was that electrons moved around the nucleus in a fixed path, which are referred to as **energy shells** or **energy levels**. Each shell has a particular size and energy level and is located at a certain distance from the nucleus.

The model was developed as a way to explain the **emission spectra** that were being produced. The interaction between an atom and energy results in light being emitted, which can be seen as distinctly coloured lines on the emission spectrum in figure 9. The emission spectrum is generated after separating the emitted light into different wavelengths by passing it through a prism.

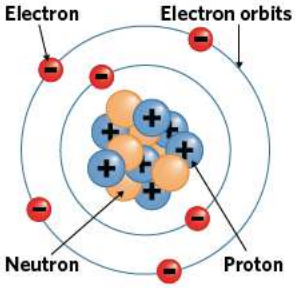


Figure 8 Bohr model of a nitrogen atom.

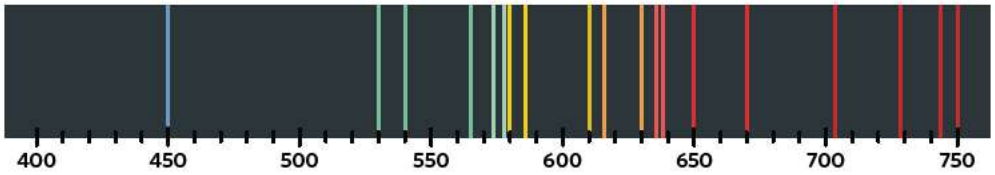
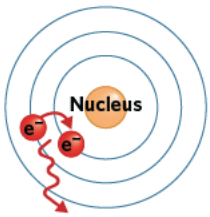


Figure 9 Example of an emission spectrum.

Bohr concluded from the emission spectrum that electrons move from the lowest energy level termed the **ground state** to a higher energy level termed the **excited state** upon receiving energy from an external source such as a flame. Electrons are most stable when they occupy a lower energy level and become less stable as they occupy higher energy levels. Since an electron is not very stable in the excited state, it returns to the ground state and in doing so emits light energy which the emission spectrum detects as coloured lines. Described in another way, when electrons get 'excited', they jump to another energy shell with a higher level of energy. When the electron does so, however, it becomes less stable and therefore returns to the original energy shell and emits energy in the form of light. This is illustrated in figure 10.



Energy released when an electron returns to its ground state. This can be seen as light.

Figure 10 Electron transitioning from its excited state to its ground state.

The presence of specific bands of light placed evenly apart indicated to Bohr that there are fixed energy shells that electrons occupy and when electrons absorb energy, they transition to an excited state. This theory allowed scientists to develop a representation to help us understand the way in which electrons are arranged in their shells. Bohr electron shell diagrams are a way of representing the arrangement of electrons, which shows electrons orbiting the nucleus in circular orbits in a similar fashion to how planets orbit the sun. This can be seen in figure 11.

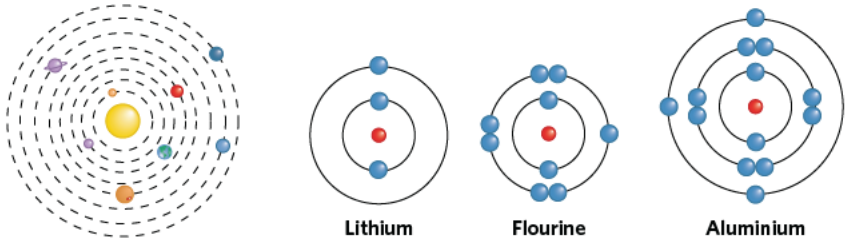


Figure 11 Electron shell diagrams of lithium, fluorine and aluminium compared to the orbiting of planets around the sun.

According to the Bohr model, the 1st, 2nd, 3rd and 4th shells are labelled K,L,M and N respectively. Table 3 summarises the key differences between the shells which differ in energy levels, distance from the nucleus and the maximum number of electrons that each shell can hold.

Table 3 Difference between shells in the Bohr model.

Name of shell	Shell number (n)	Maximum number of electrons ($2n^2$) where n is the shell number
K	1	2
L	2	8
M	3	18
N	4	32

While the Bohr model made sense of the evidence available at the time, the findings of more recent experiments have either directly contradicted the Bohr model or cannot be explained by the Bohr model. These include but are not limited to:

- The fact that the Bohr model can only explain the emission spectra of single-electron atoms.
- The fact that the Bohr model fails to explain why some emission spectra contain very fine lines.
- The fact that the Bohr model is limited to two dimensions.
- The fact that the Bohr model treats electrons as particles despite strong evidence that electrons possess wave like properties.

These limitations and new developments in physics led to the Bohr model being refined and replaced with the Schrödinger model of the atom.

Electron configurations using the shell model

We can use the Bohr model of the atom to write the **electron configurations** of the shells of different atoms. The electron configurations show information about the number of electrons in each atom and where they are located in terms of electron shells.

When electrons occupy shells in an atom, there are three main rules that need to be followed:

1 Each electron shell holds a different number of electrons.

2 Electrons will fill shells closest to the nucleus first.

3 Electrons fill shells in a particular order.

As we can appreciate from table 3, each electron shell is able to hold a particular number of electrons. We need to be aware of this number when identifying electron configurations. Electrons also prefer to fill shells with lower energy levels as these are more stable.

Consider a carbon atom, which contains six electrons. We know that the shell closest to the nucleus holds a maximum of two electrons. Therefore, carbon has its first two electrons in the lowest energy shell. This leaves us with a total of four electrons remaining and we know that the second shell can hold a maximum of eight electrons. Hence, the four remaining electrons can be added to the second shell. Therefore, the electron configuration of carbon can be written as 2,4. Figure 12 summarises this process.

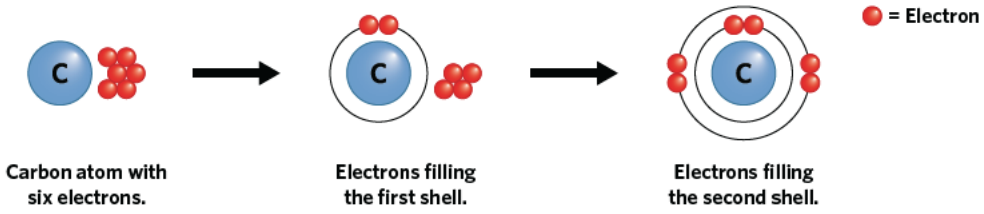


Figure 12 Carbon atom and its electrons.

An important variation to these rules occurs when electrons begin to fill the third shell. As highlighted previously, the third shell can hold up to 18 electrons. However, only eight electrons will fill the third shell after which electrons will begin to fill the fourth shell. Once the fourth shell contains a total of two electrons, the electrons will then proceed to fill the third shell again, until it reaches a total of 18 electrons.

Let's consider titanium, an atom that contains a total of 22 electrons. Once the first shell has reached its capacity, electrons begin to fill the second shell and then the third shell. This process is summarised in figure 13.

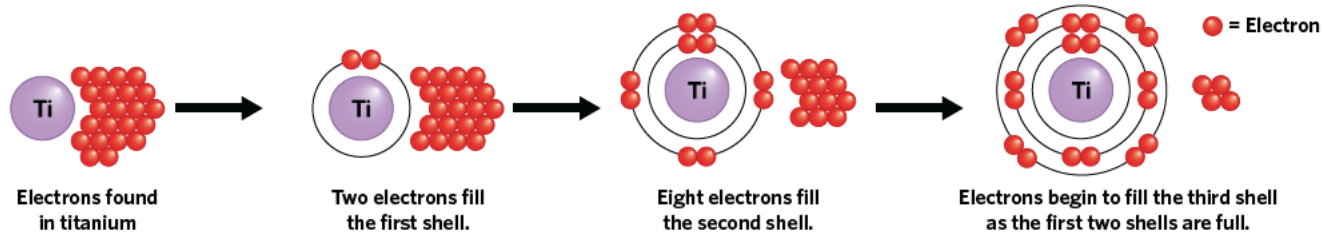


Figure 13 The filling of the first three shells of titanium.

We would think that the remaining electrons would all fill the third shell as it has the capacity to hold 18 electrons. However when the third shell reaches eight electrons, electrons begin to fill the fourth shell until the fourth shell contains two electrons. At this point, any remaining electrons will fill the third shell until it reaches its capacity of 18 as shown in figure 14.

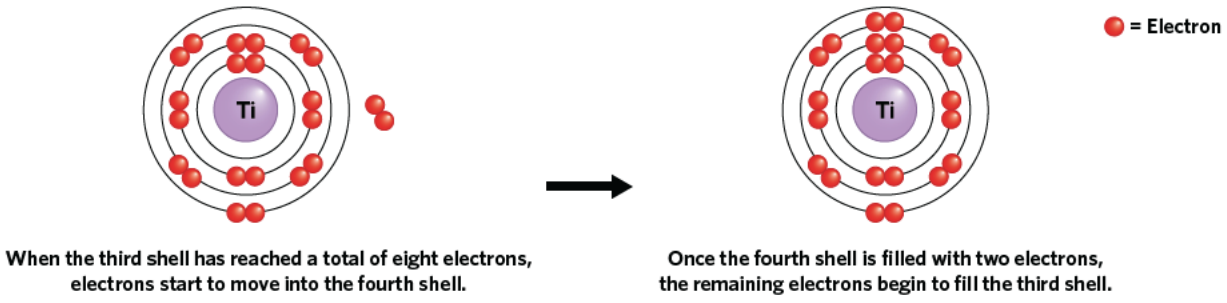


Figure 14 Electron filling of the third and fourth shell of titanium.

As a result of this process, we can see that the electron configuration of titanium is 2,8,10,2. This exception is further explored in the next section in Schrödinger's model of the atom.

Worked example 2

Scientists discovered an atom with a total of 28 electrons. What is the electron shell configuration for this atom?

What information is presented in the question?

The atom contains 28 electrons.

What is the question asking us to do?

Write the electron configuration for the atom.

What strategy(ies) do we need in order to answer the question?

1. Recall the maximum number of electrons that can fill each shell.
2. Fill each shell with electrons until it is full.
3. Write the electron configuration for the atom.

Answer

We know that the capacity of each shell is:

Shell number	Max number of electrons it can hold
1	2
2	8
3	18
4	32

However, once the third shell reaches eight electrons, it will fill the fourth shell with two electrons first before electrons will start to fill shell number three again.

According to the question, there are 28 electrons in total. Therefore the order of filling would be:

Order of filling	Number of electrons filling the shell	Total electrons placed in a shell
First shell	2	2
Second shell	8	10
Third shell	8	18
Fourth shell	2	20
Third shell	8	28

We can see that the third shell is filled twice, with a total of 16 electrons. Therefore the electronic configuration for this atom is 2,8,16,2.

The Schrödinger model 1.1.3.2

OVERVIEW

The Schrödinger model of the atom is a more accurate and complex model than the Bohr model. Instead of electrons occupying fixed orbits, electrons occupy regions around the nucleus.



This idea leads to the classic joke by Richard Feynmann that 'If you think you understand quantum mechanics, you don't understand quantum mechanics'.

THEORY DETAILS

The Schrödinger model of the atom emerged after new discoveries were made in a field of physics called **quantum mechanics**. Quantum mechanics describes the behaviour of very small objects, such as the electron. When objects are really small and have a light mass, they start to exhibit different properties to large objects which obey previously derived laws (e.g. Newton's laws). At the quantum scale, particles can start to exhibit wave-like properties and there is increased uncertainty about the speed and position of a particle. This is known as the uncertainty principle which states that the more information that is known about the velocity of a particle, the less information that is known about its position and vice versa.

Schrödinger's model of the atom takes this principle into account and defines **orbitals** as regions of space where electrons are likely to be found. It is important to note that orbitals only describe the probability of finding an electron in a particular area and are created by solving a mathematical equation known as Schödinger's equation. Some visualisations are shown in figure 15.

One of the key rules about orbitals is the **Pauli exclusion principle** which means that the maximum number of electrons in an orbital is 2. Therefore, it could hold 0, 1 or 2 electrons. However, similar to the Bohr model, electrons are said to be found in shells which vary in energy level. These shells are numbered starting at 1, and also reflect the order of energy levels. Shells contain smaller subshells, however, not all shells contain the same type of subshells as seen in figure 16.

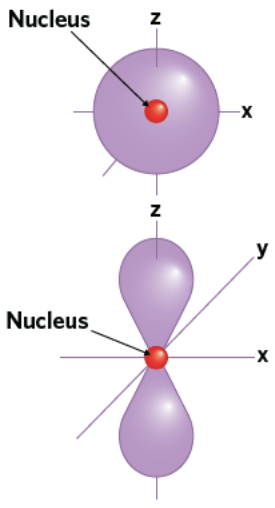


Figure 15 Different visualisation of orbitals.

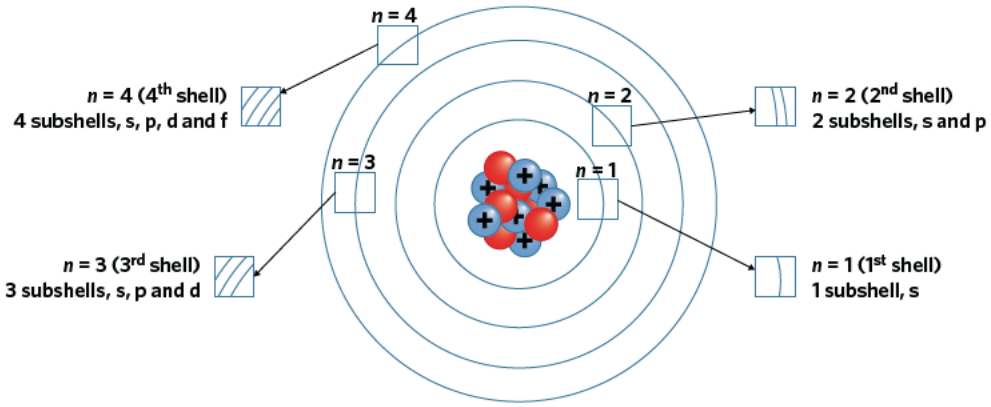


Figure 16 Shells and subshells of atoms according to the Schrödinger model.

As we can observe, an increase in the shell number also involves an increase in the number of subshells present in the shell. Each subshell has a different label, and as we can see there are the subshells s, p, d and f. The main distinguishing factor between the different types of subshells is the number of electrons they can hold, and this is a reflection of the number of orbitals found in each. These concepts are all summarised in table 4.

Table 4 Table with the different shells, subshells and orbitals in an atom.

Number of shell	Subshell(s) in shell	Number of orbitals in subshell	Maximum number of electrons
1	1s	1	2
2	2s	1	2
	2p	3	6
3	3s	1	2
	3p	3	6
	3d	5	10
4	4s	1	2
	4p	3	6
	4d	5	10
	4f	7	14



Sometimes the subshell energy levels are not in sequential order. For example, the 4s subshell is lower in energy than the 3d subshell.

Electron configurations using the Schrödinger model

We can use the Schrodinger model of the atom to also write the electron configurations of different elements by counting how many electrons are in each subshell.

In order to do this, we need to understand the **Aufbau principle** which states that electrons are fill the lowest energy level first. Based on experiments, it's been shown that the energy levels found in atoms can be represented in figure 17.

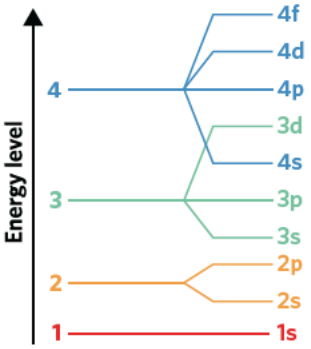


Figure 17 Different energy levels found in atoms.

While each energy level represents an energy shell, we can see that the order does not necessarily show that electrons fill shells completely before moving onto the next. Electrons prefer to fill subshells with lower energy first, as they are more stable. As such, we can see from figure 18 that electrons will begin to fill subshell 1s first, followed by 2s, 2p and so on. However, when electrons have filled subshell 3p, they then start to fill 4s before 3d. This is due to the fact that shell 4s has lower energy than 3d. This whole process can be represented by the geometric design in figure 18.

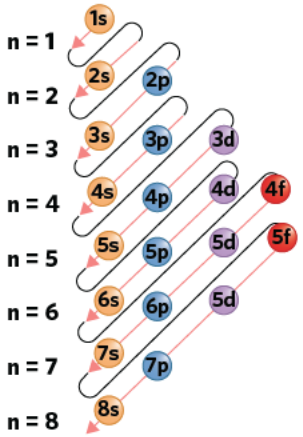


Figure 18 Pattern for remembering the order of subshell energy level.

The standard notation when writing electron configurations is to write the shell number first, the type of subshell second and then in superscript above the subshell, the number of electrons.

Consider the atom sulfur, which contains a total of 16 electrons. We know that electrons fill subshells as outlined in figure 18. We also need to remember that the s subshell is able to hold two electrons, the p subshell is able to hold six electrons and the d subshell is able to hold 10 electrons. Every subshell needs to be filled with electrons before moving onto the next subshell. Using this information, we can determine the electron configuration for sulfur as:

$1s^2 2s^2 2p^6 3s^2 3p^4$

Based on the electron configuration, we can see that sulfur contains a total of three energy shells.

Worked example 3

Write out the electron configuration according to the Schrödinger model for an atom of oxygen which has eight electrons.

What information is presented in the question?

The atom in question is oxygen which has eight electrons.

What is the question asking us to do?

Find the electron configuration of oxygen by using the Schrödinger model.

What strategy(ies) do we need in order to answer the question?

1. Write out the order of the energy levels of shells and subshells.
2. Fill the subshells with the maximum number of electrons they can hold in this order, remembering the s subshell can hold two electrons, the p subshell can hold six electrons and the d subshell can hold 10 electrons.

Answer

There are eight electrons in an oxygen atom.

The order of energy subshells is:

1s<2s<2p<3s....

The s-subshell can hold two electrons and the p-subshell can hold six electrons.

Therefore, there will be two electrons in the 1s subshell, two electrons in the 2s subshell and the remaining four electrons in the p-subshell. Therefore, the electron configuration of oxygen is:

$1s^2 2s^2 2p^4$

However, there are two common exceptions to the rule of electron filling that we need to know about, copper and chromium.

Exceptions

If we were to follow the rules exactly as listed above, we would expect that the electron configurations of copper and chromium to be:

Chromium (atomic number 24): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$

Copper (atomic number 29): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$

However, in reality the electron configurations are different for both these atoms. For chromium, the actual electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ because electrons are more stable when they are unpaired in the 4s subshell. Hence, it is energetically favourable for the 4s subshell to contain one electron and for the 3d subshell to contain five electrons. For copper, similar reasoning applies. The actual electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$. In this case, the 3d subshell is completely filled to ensure that there is a single, unpaired electron in the 4s subshell.

Theory summary

- The scale of objects ranges from kilometre to nanoscale measurements.
- The nanoscale involves particles and materials with a size around the 10^{-9} m region.
- The model of the atom has changed over time and inside an atom there are protons, neutrons, and electrons.
- The Bohr model adequately explains the emission spectrum of a single electron atom like hydrogen but is inaccurate for larger atoms.
- Schrödinger's model of the atom replaced the Bohr model and describes electrons inhabiting orbitals as well as the shells and subshells of atoms.
- The two exceptions to Schrödinger's model are copper and chromium.

2A QUESTIONS

Theory review questions

Question 1

Scale is an important concept in chemistry because it

A allows scientists to measure the relative sizes of particles.

B allows scientists to see really small particles.

Question 2

What is the multiplying factor for the nano prefix?

A 10^{9}

B 10^{-9}

Question 3

Ever since Democritus first described the atom in 400 B.C., the model of the atom has remained unchanged.

A True

B False

Question 4

_____ are positively charged and are in the nucleus of an atom whereas _____ are negatively charged and exist outside the nucleus. _____ are neutrally charged particles and, along with protons, are called _____.

A Protons, electrons, Neutrons, nucleons

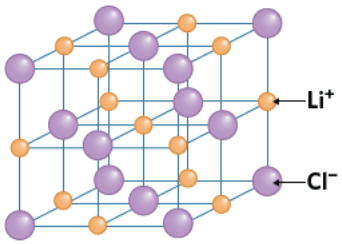
B Protons, electrons, Nucleons, neutrons

Question 5

What does the following image represent?

A A lattice

B A single atom



Question 6

What is the property of nanoparticles that allows them to function on cancer cells in the body?

A Nanoparticles are small enough to operate at the cellular level.

B Nanoparticles are chemically unreactive due to their size.

Question 7

An atom that contains 11 electrons would contain

A 3 energy shells.

B 11 energy shells.

Question 8

Which scientific observation best describes why the Bohr model was replaced with the Schrödinger model?

A The emission spectrum of atoms showed more lines than were predicted by the Bohr model.

B Atomic orbitals were seen.

Question 9

Orbitals are

A mathematical visualisations of regions where electrons are likely to be found.

B real and measurable bubbles around the nucleus.

Question 10

Which of the following are the exceptions to the Schrödinger model?

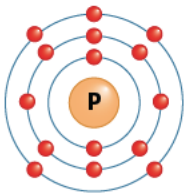
A Cobalt and titanium

B Copper and chromium

Deconstructed questions

Use the following information to answer questions 11–13.

Jasmine, a chemistry student, wanted to determine the identity of an atom based on its electron shell configuration. She was given the following diagram:



Question 11 (1 MARK)

Electron dot diagrams are based on

A the Bohr model.

B the Schrödinger model.

C the Rutherford model.

D the Plum Pudding model.

Question 12 (1 MARK)

If this is a neutrally charged atom, how many electrons are present in the atom?

A 10

B 12

C 14

D 15

Question 13 (2 MARKS)

Jasmine discovers that the atom that has 15 electrons is phosphorus. Write out the Schrödinger electron configuration for a phosphorus atom.

Exam-style questions

Within lesson

Question 14 (1 MARK)

Which of the following equations is always correct with respect to a neutrally charged atom?

A Protons = Electrons

B Protons + Neutrons = Electrons

C Protons + Electrons = 2 × Neutrons

D Electrons + Neutrons = 2 × Neutrons

Question 15 (1 MARK)

What shell and subshell will the 28th electron of an atom fill?

A 4p
B 3d
C 3s
D 3p

Question 16 (1 MARK)

Different atoms contain different numbers of electrons. Which of the following electron configurations is not possible according to the Bohr model?

	Shell number	Shell name	Number of electrons
A	4	-	31
B	-	L	19
C	3	-	15
D	-	K	1

Multiple lessons

Question 17 (10 MARKS)

Schrödinger helped to establish a different method to indicate the arrangement of electrons found in atoms.

a For each of the following atoms, write the correct electron configuration using the Schrödinger model. The number of electrons in each atom has been given.

i Lithium: 3 electrons (1 MARK)

ii Neon: 10 electrons (1 MARK)

iii Calcium: 20 electrons (1 MARK)

iv Sodium: 11 electrons (1 MARK)

v Chlorine: 17 electrons (1 MARK)

Although both models give information about the structure of an atom, the electronic configuration using the Bohr model is slightly different to that of Schrödinger.

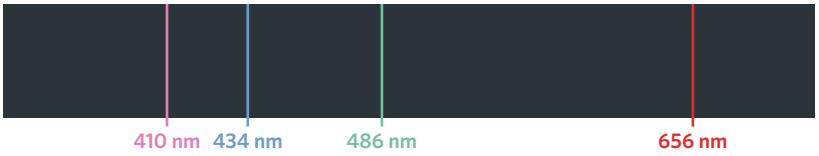
b For each of the atoms above, write the electron shell configuration based on Bohr's model. (5 MARKS)

Question 18 (9 MARKS)

The Lyman series is a set of spectral lines that is generated when the electron in a hydrogen atom transitions from a shell (n) where $n > 1$ to $n = 1$.

a Identify the name of the state that the electron is in immediately after it has absorbed energy from the ground state. (1 MARK)

b The following image is the emission spectrum for the hydrogen atom. If the electron in the hydrogen atom starts in shell 1, what is the maximum shell number it reaches and why? Assume each transition from excited to ground state is to the first shell. (2 MARKS)



c Bohr's model of the atom is currently not in use today. With reference to three limitations, discuss why this is the case. (3 MARKS)

d Identify the most recent model of the atom and describe how it differs to the Bohr model. (3 MARKS)

Question 19 (10 MARKS)

In the realm of quantum mechanics, there is something known as quantum numbers. Electrons are given a set of four quantum numbers which describe the size, shape and orientation the orbital is in, as well as the direction that the electron is spinning in.

a What scale of objects does quantum mechanics apply to and why might it be suitable for describing electrons? (2 MARKS)

b How does the p orbital differ to the s orbital? Give three differences. (3 MARKS)

c Jason, the chemistry student, deduced that the 3d orbital of an atom would be filled before the 4s orbital because the order of energy levels goes 1s, 2s, 2p, 3s, 3p, 3d, 4s etc. Is the student correct? Explain why or why not. (2 MARKS)

d Write the electron configuration of a copper atom and a zinc atom which contains 29 and 30 electrons respectively. What is the same about these two electron configurations that would not normally be expected? (3 MARKS)

Key science skills

Question 20 (6 MARKS)

A company regulator is investigating the effectiveness of sunscreens with nanoparticles. In order to test this, an experiment is set up with the use of infrared absorbing beads which change colour when exposed to increasing levels of radiation. Half the beads were covered with sunscreen without nanoparticles and half the beads were covered with sunscreen with nanoparticles.

a Identify the independent variable in this experiment. (1 MARK)

b Identify the controlled variables. (3 MARKS)

c Suggest a reason why nanoparticles are used in sunscreen. Hint: the goal of sunscreen is to absorb as much radiation as possible. (1 MARK)

d Health concerns have been raised about the use of nanoparticles in sunscreen. Suggest a reason why this might be the case. (1 MARK)

Exam-style question hints

14: For an atom to be neutral, charged particles need to cancel each other out. **15:** Sometimes, electrons do not completely fill shells before moving onto the next shell. **16:** Each shell has a maximum number of electrons it can hold. **17a:** Schrödinger's model of the atom includes subshells. **17b:** The electron shell configuration is written differently than the electron configuration based on Schrödinger's model **18a:** Given enough energy, electrons can transition between shells. **18b:** When electrons transition from higher energy shells, they emit energy that can be visualised. **18c:** Bohr's model can only explain the main features of the atom. **18d:** Quantum mechanics gave rise to a new model of the atom. **19a:** Quantum mechanics is used to explain the properties of very small objects. **19b:** Each orbital has its own properties. **19c:** Each orbital has its own level of energy. **19d:** Electrons fill shells in a particular order. **20a:** The independent variable is the variable that is manipulated during the experiment. **20b:** Controlled variables are those that are kept constant throughout the experiment. **20c:** Nanoparticles assist substances with absorption. **20d:** Nanoparticles are extremely small.

2B DEFINING ELEMENTS

As we learnt in the last lesson, atoms consist of three different components – protons and neutrons which form the nucleus and electrons which are located in orbitals that surround the nucleus. In this lesson we'll discuss how to accurately describe the structure of elements and learn about the characteristics that define elements.

2A The atom | 2B Defining elements | 2C The periodic table - part 1 | 2D The periodic table - part 2 | 2E Relative masses | 2F The mole - part 1 | 2G The mole - part 2 | 2H Chemical formulas and equations

Study design dot point

- the definition of an element with reference to atomic number; mass number; isotopic forms of an element using appropriate notation

Key knowledge units

Atomic number	1.1.2.1
Mass number	1.1.2.2
Isotopes	1.1.2.3

Key terms and definitions

Element pure substance containing only a single type of atom

Periodic table table of chemical elements in which elements are arranged in order of increasing atomic number

Chemical symbol abbreviation used to represent a chemical element

Molecule two or more atoms chemically bonded together

Compound two or more atoms of different elements bonded together

Atomic number number of protons in the nucleus of an atom

Mass number mass of an atom, approximately equal to the average sum of protons and neutrons

Isotope variants of an element which have the same atomic number but a different number of neutrons in their nuclei

Lesson links

This lesson builds on:

▶ 2A The atom

There are many different types of atoms that exist.

Atomic number 1.1.2.1

OVERVIEW

All elements have different atomic numbers.

THEORY DETAILS

Now that we know the basic structure of an atom, we can start to appreciate the many different atoms that exist. To date, scientists have discovered 118 different types of atoms which are referred to as **elements**. With so many different elements to keep track of, Russian chemist Dimitri Mendeleev developed a way of organising these into an organised table we now refer to as the **periodic table** shown in figure 1. We will learn about the periodic table in great detail in the next few lessons.

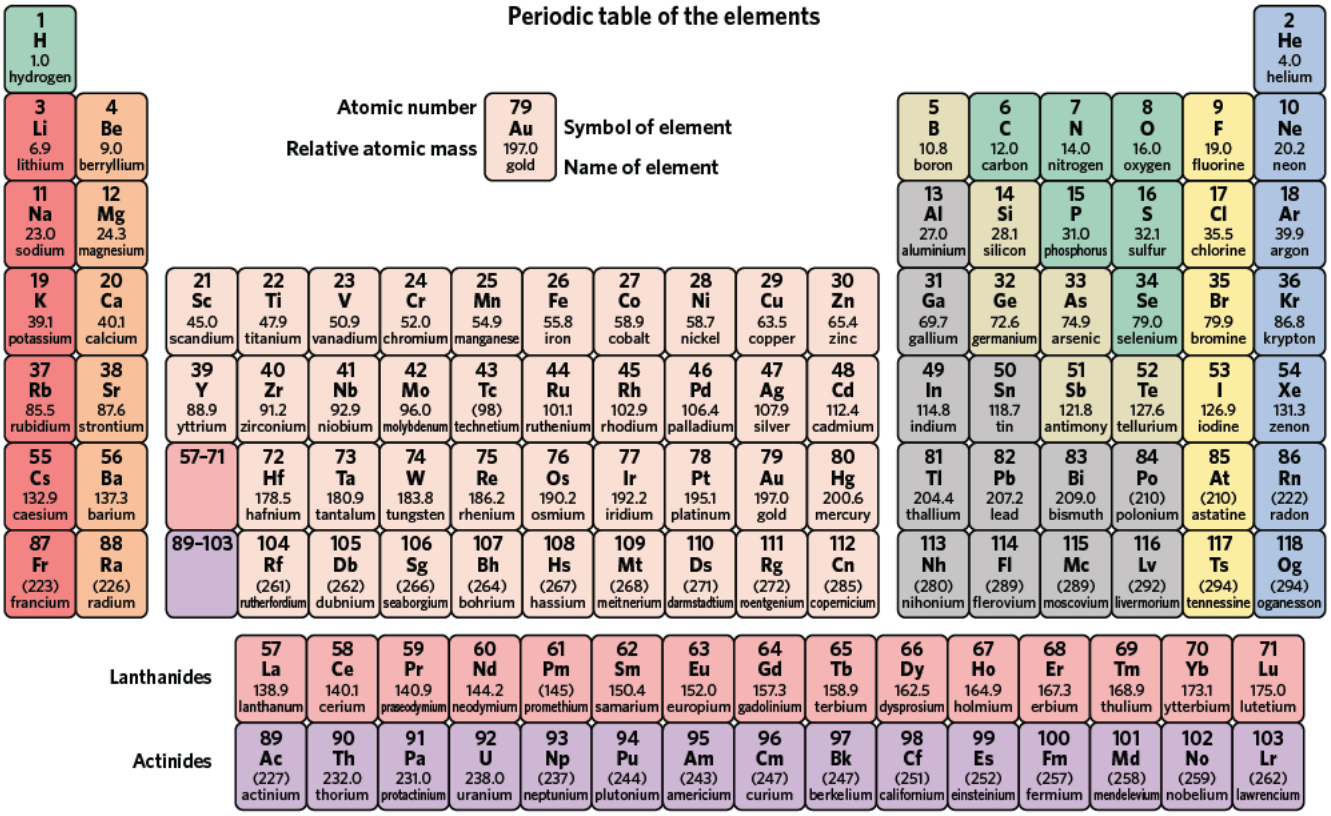


Figure 1 The periodic table.

Although we have names for each of the different elements, chemists generally represent each element using its **chemical symbol**. Each box in the table represents a new element, as seen by the different names and symbols. Generally speaking, the chemical symbol for an element consists of a letter or letters found in its actual name, however this does not always occur as in the case of mercury which has the chemical symbol Hg.

When two or more atoms bond together, they form **molecules** such as O_2 (oxygen). When these atoms come from different elements, such as H_2O (water), they can form **compounds**. By doing so, we can create many different materials that are important in everyday life.

We will now explore the information displayed for each element on the periodic table in further detail. As shown in figure 2, the number located at the top of each element indicates how many protons are in the nucleus of each atom and is referred to as the **atomic number**.

The atomic number is unique to every element and gives information about the number of protons found in the atom. Looking at the example of oxygen in figure 2, oxygen has an atomic number of 8, and by no coincidence, is also the 8th element on the periodic table. This means that the identity of an element is based on the number of protons it has in its nucleus, and does not depend on the number of neutrons or electrons.

In a neutrally charged single atom, the atomic number of an element is equivalent to the number of electrons in orbit around the nucleus. In other words, a proton has a charge of +1, and an electron has a charge of −1, so if the quantity of protons and electrons in the atom are equal then the atom will be expected to have an overall neutral charge.



Chemical compounds are like compound words – a word made by joining two smaller words together.

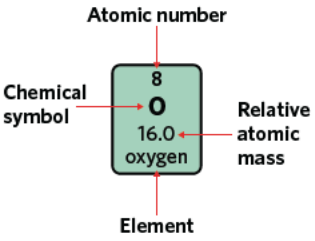


Figure 2 Basic representation of elements in the periodic table.



The identity of an element is determined by the number of protons it has in its nucleus.

Mass number 1.1.2.2

OVERVIEW

The mass number of an element indicates the number of protons and neutrons in an atom.

THEORY DETAILS

As elements are made up of a different combination of subatomic particles, they all have different masses. The **mass number** of an element is based on the number of protons and neutrons found in the nucleus and is calculated as:

Mass number = number of protons + number of neutrons

If we inspect the model of an oxygen atom in figure 3, we can see that the nucleus contains a total of eight protons and eight neutrons. To calculate the mass number for oxygen, we need to sum the number of protons and neutrons.

Mass number for oxygen = 8 + 8 = 16

Notice that we have not included the mass of electrons as part of our calculations. In the previous lesson, we learned that electrons have a very small mass; so small that it's almost negligible when it comes to calculating the mass of an atom. As a result, we don't include electrons in our calculations.

If we combine our knowledge of the atomic number and mass number of an element of an atom, we can determine the number of neutrons found in an atom using the following formula:

Number of neutrons in an atom = mass number – atomic number

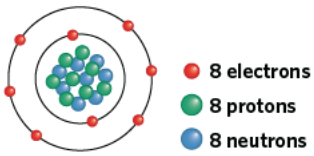


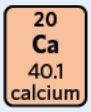
Figure 3 Bohr model of an oxygen atom.



The mass number doesn't include electrons.

Worked example 1

Consider the following element. Given that the mass number is 40, how many neutrons would be found in the nucleus of this atom?



What information is presented in the question?

Information from the periodic table for the element calcium.

What is the question asking us to do?

Determine the number of neutrons found in the nucleus of calcium.

What strategy(ies) do we need in order to answer the question?

1. Determine the atomic number of calcium.
2. Subtract the atomic number from the mass number.

Answer

The atomic number of calcium is 20. Given that it has a mass number of 40, the number of neutrons would be:

Neutrons = 40 – 20 = 20

There would be 20 neutrons in a calcium atom.

Isotopes 1.1.2.3

OVERVIEW

There are different variations of the same element.

THEORY DETAILS

Individual elements can exist in different forms which are referred to as **isotopes**. Isotopes of an element contain the same number of protons (and therefore have the same atomic number) however vary in the number of neutrons. As we can see from figure 4, hydrogen exists in many different forms all of which differ by the number of neutrons. The isotopes are referred to as protium, which contains only a single proton in its nucleus, deuterium, which contains one proton and one neutron in the nucleus and tritium, which contains two neutrons and one proton in the nucleus. Since all three contain only one proton in their nucleus, the atomic number for all isotopes is one. The difference in the number of neutrons affects the mass number and behaviour of the hydrogen isotopes.

Considering the mass number is the total number of protons and neutrons in the atom, the mass numbers of protium, deuterium and tritium would be 1, 2 and 3 respectively. The mass number is shown as a superscript to the upper left hand side of the chemical symbol. This value can be used to determine the identity of the isotope as shown in figure 5.

The isotopes of an element do not exist in the universe at the same proportions. For example, protium is by far the most abundant isotope of hydrogen and accounts for about 99.99% of hydrogen atoms found in nature. The addition of neutrons to the nucleus of an atom can cause the atom to become unstable and even radioactive resulting in the emission of radiation.

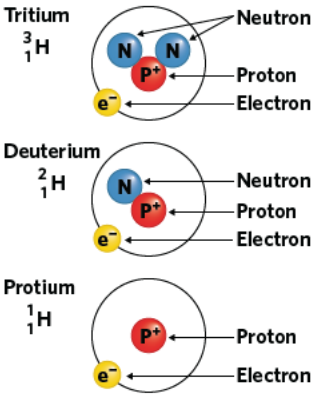


Figure 4 Different isotopes of hydrogen.

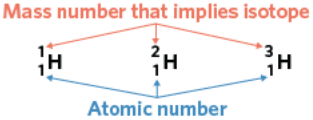


Figure 5 Conventions used to represent isotopes.



Isotopes are like siblings who share the same core characteristics yet also possess distinctive, unique characteristics.

Theory summary

- The atomic number reflects the number of protons in the nucleus of an atom.
- Different elements have different atomic numbers.
- The mass number is the sum of neutrons and protons in the nucleus of an atom.
- Isotopes have the same number of protons but a different number of neutrons.

2B QUESTIONS

Theory review questions

Question 1

The atomic number describes

A the number of nucleons in an atom.

B the number of protons in an atom.

Question 2

The mass number is determined by

A multiplying the mass of a proton by the number of protons.

B adding the total number of neutrons and protons in the nucleus of an atom.

Question 3

A neutrally charged atom with 10 electrons would definitely have

A 10 neutrons.

B 10 protons.

Question 4

A neutrally charged atom with 10 protons would definitely have

A 10 neutrons.

B an atomic number of 10.

Question 5

Elements on the periodic table are numbered in increasing order of

A the number of protons.

B the number of electrons.

Question 6

A proton weighs approximately the same amount as a

A neutron.

B electron.

Question 7

The number of neutrons in an atom with an atomic number of 11 and a mass number of 23 is

A 12.

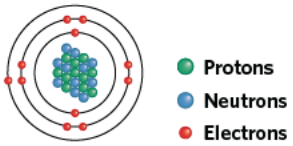
B 34.

Question 8

The mass number of the following atom is

A 22.

B 44.



Question 9

Isotopes have

A atoms with a different number of protons but the same number of neutrons.

B atoms with a different number of neutrons but the same number of protons.

Question 10

Isotopes of an element all behave the same.

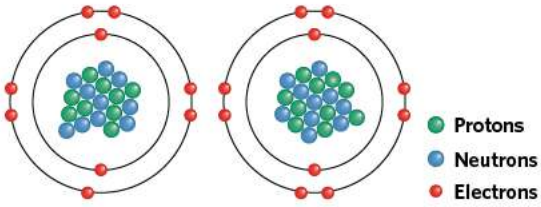
A True

B False

Deconstructed questions

Use the following information to answer questions 11–13.

Atoms are the most basic building blocks of living things. In order to be able to produce different kinds of substances, we require a whole range of different types of atoms. Take for example the following two atoms:



Question 11 (1 MARK)

These atoms

A are isotopes of each other.

B have the same number of neutrons.

C have the same number of protons.

D are radioactive.

Question 12 (1 MARK)

The atomic number of each atom is

A 19 and 20 respectively.

B 10.

C 9 and 10 respectively.

D 7 and 8 respectively.

Question 13 (2 MARKS)

Describe the difference in mass number between the two atoms.

Exam-style questions

Within lesson

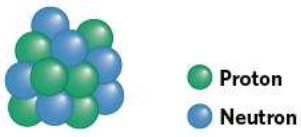
Question 14 (1 MARK)

Dorothy is a nuclear physicist who devotes much of her time to her love of isotopes. One afternoon she decides that she wishes to explore the potential of artificially synthesising some zinc isotopes, and carries out some inquiries. Identify the key characteristic of zinc that she would change in order to develop different isotopes.

Question 15 (2 MARKS)

When an atom becomes radioactive, it can undergo a process called neutron release. During neutron release, a neutron is effectively 'ejected' from the nucleus.

Consider an atom with the nucleus shown. Assuming all of the nucleons are visible, describe what would happen to the atom if it were to undergo neutron release.



Question 16 (8 MARKS)

Alicia conducted an experiment to detect subatomic particles in newly discovered atoms. Her observations are given in the table below:

Elemental symbol	Protons	Neutrons	Electrons	Mass number
R	17	17	17	
I	18	18	18	
T	18	19	18	
A	19	20	19	

a Determine the mass number of each element. (4 MARKS)

b Which of the two elements are isotopes of each other? (1 MARK)

c How should the table be updated to reflect your answer in part b? (1 MARK)

d Fill in the gaps below to represent element A. (1 MARK)

__—
__—

e If elements R and A came together to create RA, would RA be classified as a molecule or a compound? (1 MARK)

Multiple lessons

Question 17 (3 MARKS)

A university professor has been asked to construct a table showing the distribution of isotopes of bromine. It is known that a neutral bromine atom has an atomic number of 35.

The professor first wishes to establish some fundamental information, and obtains a neutrally charged isotope of bromine with a mass number of 81.

a How many protons, neutrons and electrons are there in the isotope? Explain your reasoning. (2 MARKS)

b Write the electron configuration for the isotope of bromine. (1 MARK)

Question 18 (3 MARKS)

A neutral element with the electron configuration of $1s^2 2s^2 2p^6 3s^1$ was found to be metallic in nature. It was also found that the element had a mass number of 23.

a Identify the number of energy shells found in the atom. (1 MARK)

b Identify the atomic number of this element. (1 MARK)

c Identify the number of nucleons found in the atom. (1 MARK)

Key science skills

Question 19 (3 MARKS)

Matthew, a Year 11 chemistry student, wanted to compare two unknown elements. He believed that by developing an experiment which could measure and compare the mass number of two elements, that he would be able to distinguish the two elements from one another.

In preparation for his experiment, Matthew developed a set of scales that could detect extremely small masses. Prior to conducting the experiments, he made sure to calibrate the scales.

a Why is it important for Matthew to calibrate the scales prior to an experiment? (1 MARK)

Consider the following statement from Matthew:

'Every element is different, and therefore will have a different mass number to other elements.'

b Comment on the accuracy of his statement. (2 MARKS)

Exam-style question hints

14: Isotopes have a varying number of neutrons. **15:** Isotopes have a varying number of neutrons. **16a:** Calculating mass number involves only the nucleons. **16b:** Isotopes are derived from the same element. **16c:** Isotopes have the same atomic number. **16d:** The representation involves only symbols and numbers. **16e:** Compounds involve two or more different elements. **17a:** The mass number can help to determine the number of nucleons. **17b:** Energy shells contain subshells. **18a:** The electron configuration indicates the number of energy shells. **18b:** Neutral atoms have equal numbers of electrons and protons. **18c:** Mass number reflects the number of protons and neutrons. **19a:** Errors can affect experimental results. **19b:** Isotopes have different masses.

2C THE PERIODIC TABLE - PART 1

In this lesson, we will learn about how elements are organised in the periodic table.
More specifically, with regards to their atomic radius, ionisation energy and electronegativity.

2A The atom	2B Defining elements	2C The periodic table - part 1	2D The periodic table - part 2	2E Relative masses	2F The mole - part 1	2G The mole - part 2	2H Chemical formulas and equations

Study design dot point

- the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic/non-metallic character and reactivity) of elements

Key knowledge units

The basic organisation of the periodic table	1.1.4.1
First ionisation energy	1.1.4.3
Electronegativity	1.1.4.2

Key terms and definitions

Periods rows in the periodic table

Periodicity characteristics of elements in a period

Atomic radius size of an atom measured as the distance from the nucleus to the valence shell in picometers (pm)

Valence shell outermost energy shell

Valence electrons electrons in the outermost shell of an atom

Groups columns in the periodic table

Ion charged atom

First ionisation energy energy required to remove the first valence electron from an atom

Core charge attractive force between the positive nucleus and valence electrons, measured as the difference in the number of protons and inner electrons

Electronegativity how strongly an atom attracts electrons towards itself

Lesson links

This lesson builds on:

- 2B Defining elements

The periodic table groups atoms based on their properties.

The basic organisation of the periodic table 1.1.4.1

OVERVIEW

The periodic table is organised in a systematic manner.

THEORY DETAILS

As we explored in the previous lesson, an element is a type of atom. Elements behave differently, and the periodic table helps scientists represent all the different elements in a way that demonstrates their properties and trends. A modern version of the periodic table is shown in figure 1.

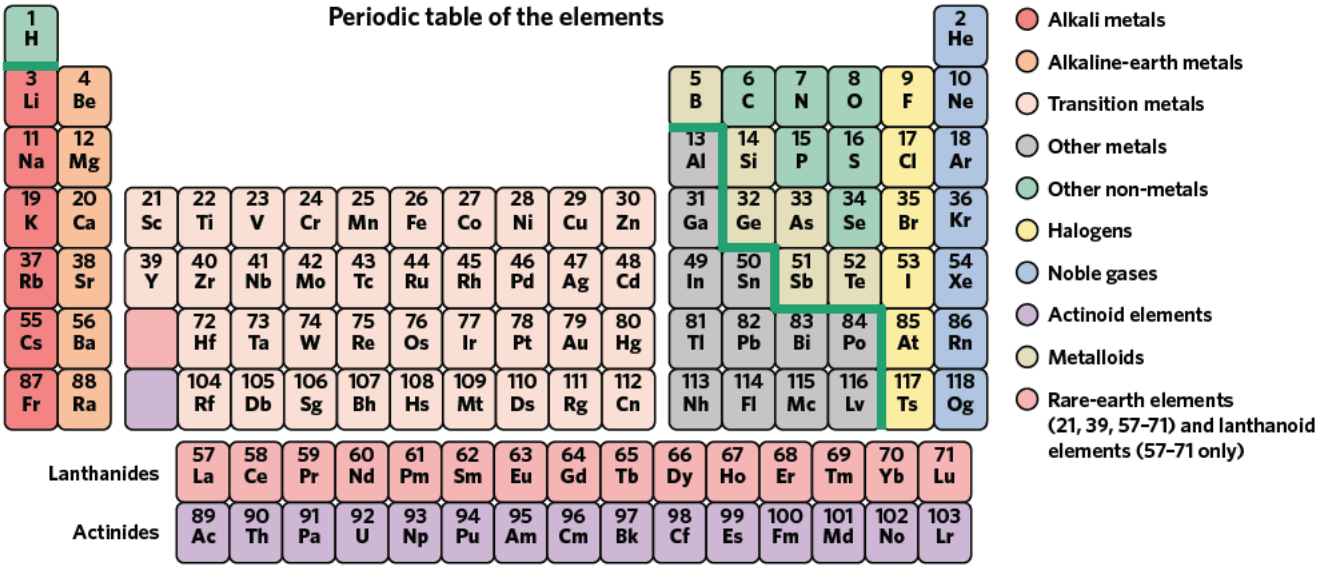


Figure 1 A modern periodic table.

From a broad perspective, we can see that the colours in the periodic table highlight the different types of elements that exist; ranging from metals to nonmetals. Within these two categories, the elements can be further classified into sub-categories.

Periods

At first glance, we can notice that elements are arranged by their atomic number, which increases from left to right across rows. As we learned in the previous lesson, the atomic number reflects the number of protons in the nucleus of an atom, therefore we can appreciate that as we progress from left to right on the periodic table, the number of protons increases.

Rows in the periodic table are referred to as **periods.** Elements in the same period show **periodicity;** a repeating trend in properties, including:

- Same number of electron shells
- Decreasing in atomic size

Electron shells of elements in the same period

A special characteristic of periods on the periodic table is that elements in the same period contain the same number of electron shells as seen in figure 2.

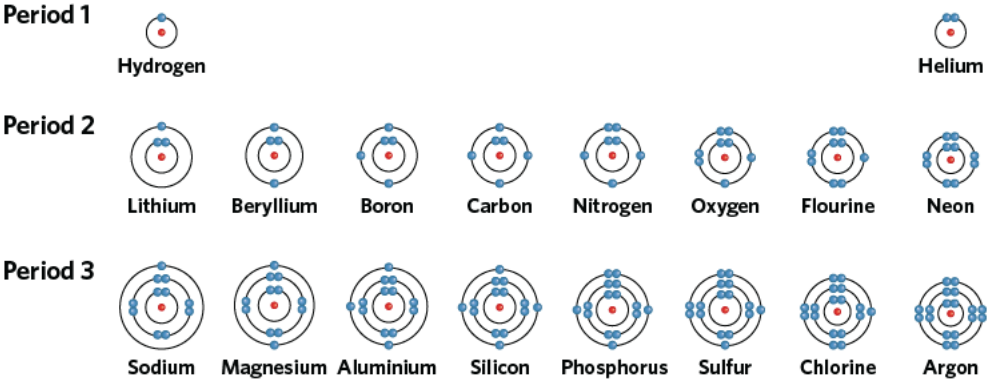


Figure 2 The electron arrangement in different elements.

As we learned in lesson 2A, elements contain different numbers of electrons and therefore varying numbers of electron shells. Take period number one, which contains two elements: hydrogen and helium. As we can see, both hydrogen (the first element on the periodic table) and helium (the second element on the periodic table) contain only one electron shell. According to our understanding of electron configurations, we know that the first energy shell can only hold a maximum of two electrons. As such, when moving to lithium (the third element on the periodic table), we can see that it occurs on the next period on the periodic table, period two, and thus contains two electron shells in its atomic structure. If we continue to move across period two, we can appreciate that the elements all contain two electron shells. This trend continues throughout the rest of the periodic table.

Atomic radius of elements in the same period

If we look carefully at figure 2, we can also see that although elements in the same period contain the same number of electron shells, the size of the atom actually becomes smaller. This seems counterintuitive considering as we move from left to right across the period, there is an increase in the number of protons and therefore also electrons. However, if we investigate further, we will realise that the increase in protons is exactly the reason why the size of the atom decreases as we move along a period.

The **atomic radius** of an atom is measured from the nucleus to the outer shell of an atom as shown in figure 3. The outermost shell of an atom is known as the **valence shell**. Electrons that are found in the valence shell are known as **valence electrons.** Valence electrons are an important contributor of the behaviour of an atom; a concept that we will investigate further in this lesson and the next lesson.

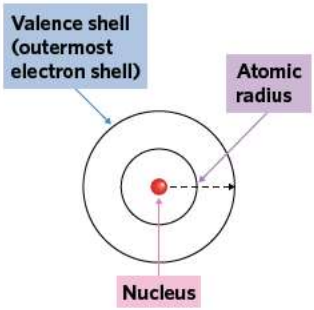


Figure 3 The method of measuring atomic radius.

As mentioned in lesson 2A, electrons are held to the atom due its attraction to the positively charged protons in the nucleus. Considering the number of protons in an atom increases with atomic number across a period, there will be more protons in the nucleus and therefore there would be a greater positive charge in the nucleus. Together with the fact that each period contains the same number of shells, atoms towards the end of the period would have a greater ability to attract the valence electrons towards the nucleus compared to elements at the start of the period, due to the increase in number of protons (and therefore positive charge). This would result in the valence shell being pulled closer to the nucleus, reducing the atomic radius. This is summarised in figure 4.

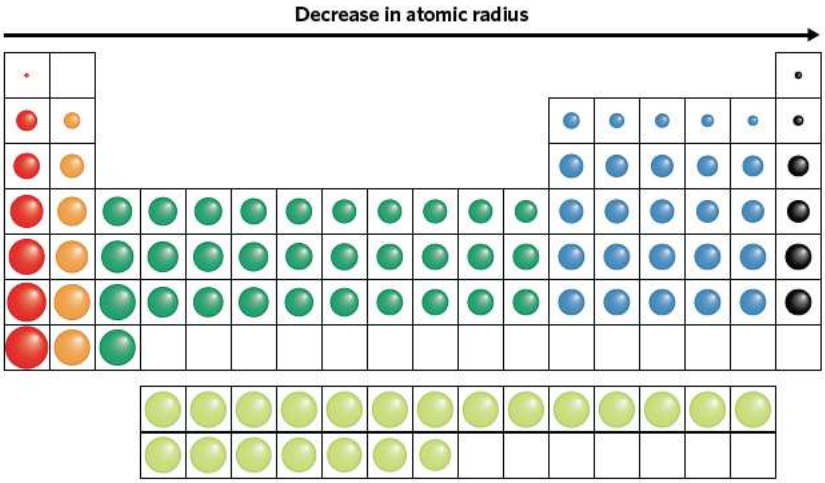


Image: magnetix/Shutterstock.com

Figure 4 Periodicity of elements in the same period.

However, we need to remember that in between the valence shell and the nucleus is the presence of electrons found in other shells. As we know, electrons are negatively charged particles, and therefore would repel each other when they come into close contact, as shown in figure 5.

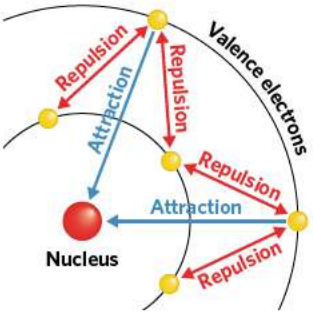


Figure 5 Repulsion of electrons in an atom.

So far, we have been investigating the periodic table by looking at the properties of elements in the same period (row) however we can also look at the table from a different perspective; investigating elements in the same column.

Groups

The columns of a periodic table are known as **groups**. Groups are numbered from left to right from 1 to 18, where each different column represents a different group. Consider figure 6, which shows the first few elements in a variety of different groups.

If we look closely, we can notice that the valence shell in each element (in red) of each group contains the same number of valence electrons. Hydrogen, lithium and sodium, all found in group 1, all contain one valence electron. The elements in group 14 all contain four valence electrons. This pattern is evident throughout all groups. Due to the fact that elements in the same group contain the same number of valence electrons, they all behave similarly.

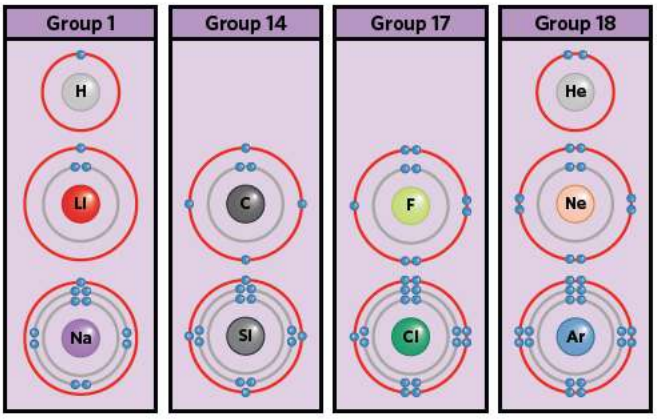


Figure 6 Electron arrangement of elements in the same group.

Atomic radius of elements in the same group

Although elements in the same group have the same number of valence electrons, they differ in the total number of electron shells. As we move down a group, we can see that there is an additional electron shell. This makes sense considering when we move down a group, we are essentially entering a new period. As such, we experience the addition of another electron shell.

The increase in the number of electron shells as we progress down a group results in an increase in atomic radius as seen in figure 7. As we know, this means that there is an increase in the distance between the nucleus and the valence shell. Consequently, the valence electrons experience a weak force of attraction towards the nucleus. Again, the electrons found in between the valence shell and nucleus also contribute to the atomic radius.

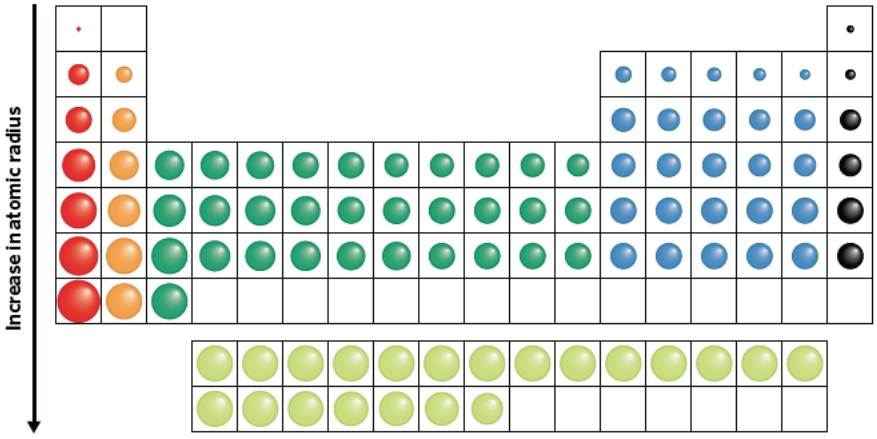


Image: magnetix/Shutterstock.com

Figure 7 The trend in atomic radius of groups in the periodic table.

First ionisation energy 1.1.4.3

OVERVIEW

The properties of atoms affects the amount of energy required to remove electrons from an atom.

THEORY DETAILS

The way in which elements are represented on the periodic table is as neutral atoms. That means that the number of electrons and protons are the same, and therefore the negative and positive charges cancel each other out. However, this is not how elements always exist.

Given enough energy, we are able to remove electrons from an atom. More specifically, atoms are able to lose their valence electrons. When this happens, atoms become **ions**; a charged atom. The amount of energy required to remove the first valence electron from an atom in its gaseous state is known as the **first ionisation energy**. This depends on the attraction of the valence electrons to the positively charged nucleus.

As we learned above, the atomic radius of an atom decreases as we move across a period and increases as we move down a group, which affects the attraction between the nucleus and the valence electrons. This concept, the attraction between the nucleus and valence electrons, is related to the **core charge** of an atom; which describes how strongly the nucleus is able to attract valence electrons. The core charge takes into consideration the electrons between the valence shell and the nucleus, also known as the inner-shell electrons, and can be mathematically expressed as:

Core charge = atomic number (number of protons) – number of inner-shell electrons.

The greater the core charge, the greater the attraction between the valence shell and the nucleus, and therefore the more energy would be required to remove valence electrons from the atom. We also need to consider however the impact of the inner shell electrons on the ability of the nucleus to attract the valence electrons. The inner electrons can shield the valence electron from feeling the full attractive force of the nucleus as seen in figure 8, and is why we need to consider these electrons when calculating the core charge.

Consider a lithium and sodium atom as shown in figure 9, both of which occur in group 1.

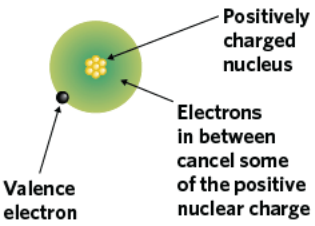
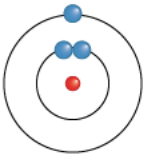
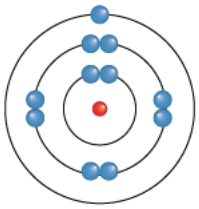


Figure 8 Illustration of the shielding effect of inner shell electrons.



Lithium

Protons: 3
Electrons: 3
Core charge = 3 – 2 = +1



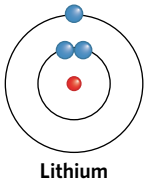
Sodium

Protons: 11
Electrons: 11
Core charge = 11 – 10 = +1

Figure 9 The deconstruction of a lithium and sodium atom.

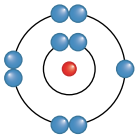
We can see that both lithium and sodium have a core charge of one, therefore we would expect the valence electrons from both atoms to experience the same level of attraction to the nucleus. However, sodium contains more inner-shell electrons and has a larger atomic radius, therefore not only would there be more shielding from inner-shell electrons, but the valence electron is much further from the nucleus compared to lithium. As a result, sodium is not able to hold onto its valence electron as strongly as lithium, and thus would require less energy to remove the valence electron. The first ionisation energy for sodium would be less than that of lithium.

We can also compare lithium with fluorine as shown in figure 10, both of which are found in period 2. As we can see, fluorine has a greater core charger and smaller atomic radius than lithium, both of which allows fluorine to better attract the valence electrons. As a result, fluorine would have a greater first ionisation energy than lithium. The overall trend of ionisation energy can be seen in figure 11.



Lithium

Protons: 3
Electrons: 3
Core charge = 3 − 2 = +1



Fluorine

Protons: 9
Electrons: 9
Core charge = 9 − 2 = +7

Figure 10 The deconstruction of a lithium and fluorine atom.

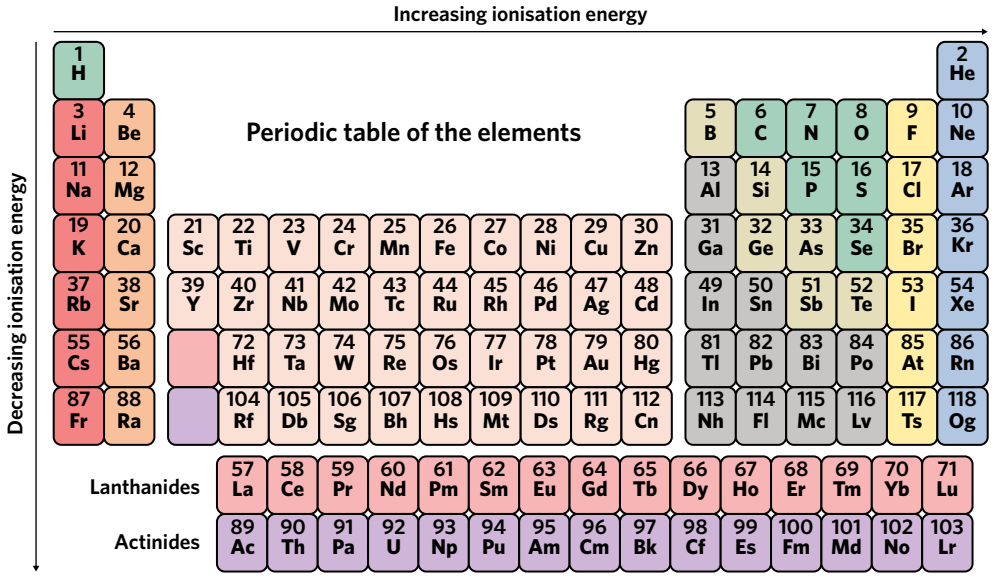


Figure 11 Trend in ionisation energy of elements in the periodic table.

We can now appreciate that both core charge and atomic radius affect the first ionisation energy of an atom.

Electronegativity 1.1.4.2

OVERVIEW

Different elements attract valence electrons differently.

THEORY DETAILS

As we learned above, different atoms have the ability to attract electrons differently depending on the core charge and the atomic radius. The ability of an atom to attract electrons to itself is also known as **electronegativity**. The greater the electronegativity, the greater the ability of the atom to attract electrons to its nucleus. This will be an important concept when we explore chemical bonding later in the book.

The quantitative measurement of electronegativity is called the Pauling Scale, where the greater the number, the more electronegative the element. As we can see in figure 12, fluorine has the highest number on the Pauling Scale, and therefore is the most electronegative element on the periodic table, whereas caesium has the lowest Pauling value, therefore is the least electronegative element on the periodic table. For this study design, we are not required to know the Pauling Scale in any detail.

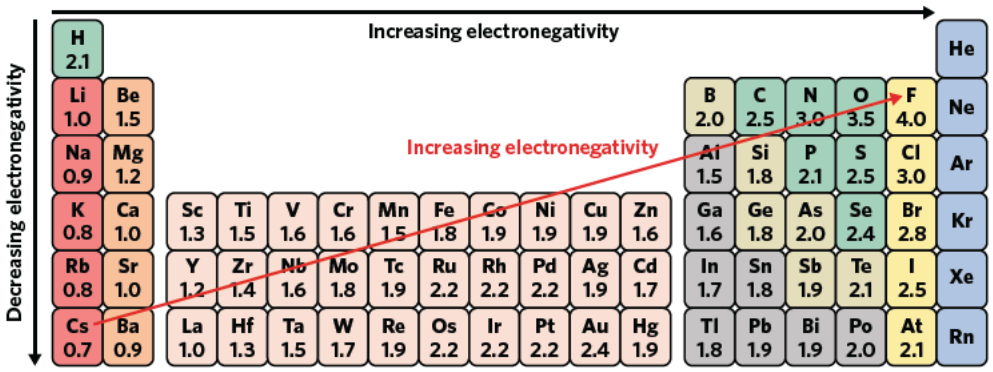
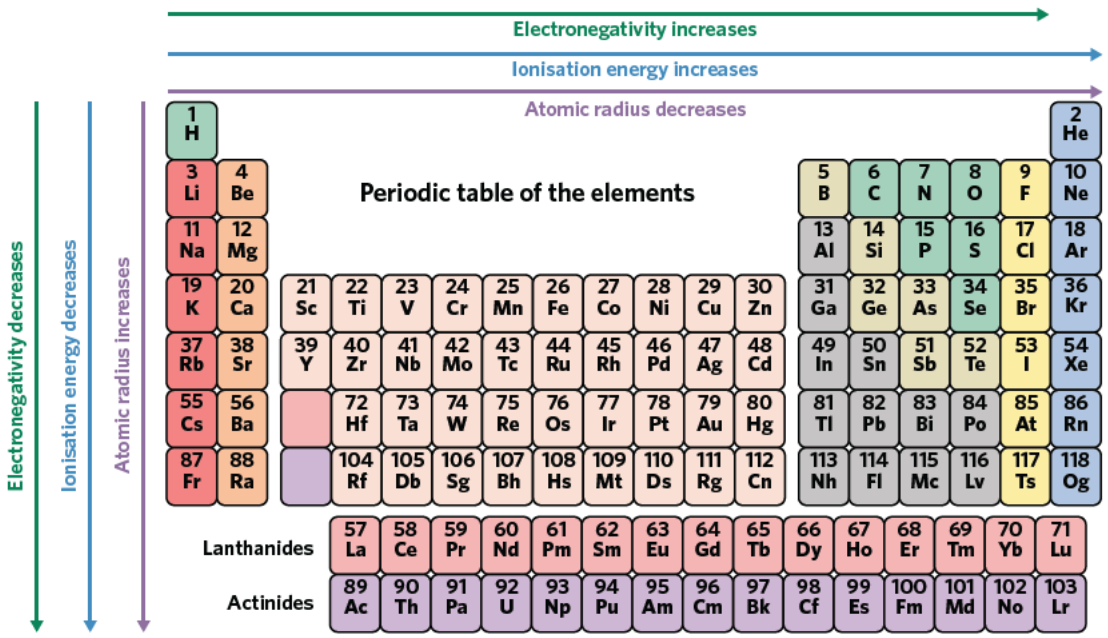


Figure 12 Trends in electronegativity in the periodic table.

The trend highlighted matches with our understanding of the core charge and atomic radius of elements in the periodic table. Elements with a higher core charge and smaller atomic radius can be found towards the end of each period and further up in a group. As such, the trend in electronegativity follows the same pattern, where the most electronegative elements are found towards the upper right hand corner of the periodic table. A notable exception is that elements in group 18 of the periodic table (inert or noble gases) have no electronegativity. This is because they already have eight electrons in their valence shell and therefore, according to the octet rule, do not want to attract any more electrons.

Theory summary



	→ a period (left to right)	↓ a group
Atomic radius	decreases	increases
Ionisation energy	increases	decreases
Electronegativity	increases	decrease
Core charge	increases	unchanged

2C QUESTIONS

Theory review questions

Question 1

Elements with the same number of energy shells are found in the same

A group.

B period.

Question 2

Elements with three valence electrons can be found in the same

A group.

B period.

Question 3

Potassium can be found in

A group 1, period 4.

B group 4, period 1.

Question 4

The core charge _____ as we move across a period.

A decreases

B increases

Question 5

The larger the atom, the greater the core charge.

A True

B False

Question 6

Atomic radius increases across a period.

A True

B False

Question 7

Atomic radius increases as core charge increases.

A True

B False

Question 8

Ionisation energy increases as we move

A down a group.

B up a group.

Question 9

Ionisation energy decreases as atomic size increases.

A True

B False

Question 10

Highly electronegative elements will

A have a large atomic radius.

B strongly attract electrons.

Deconstructed questions

Use the following information to answer questions 11–13.

Selenium is an element found in trace amounts in everyday foods. A sample of selenium is examined to determine its chemical properties. A second element is studied, sulfur, which exists abundantly on Earth as a yellow crystallised solid.



Images (left to right): Bjoern Wylezich, Sebastian Janicki/Shutterstock.com

Question 11 (1 MARK)

Sulfur is found in group

A 13.
B 14.
C 15.
D 16.

Question 12 (1 MARK)

In comparison to sulfur, selenium has

A the same number of electron shells.
B a larger core charge.
C less valence electrons.
D a larger atomic radius.

Question 13 (2 MARKS)

Which element would have the lowest ionisation energy?

Exam-style questions

Within lesson

Question 14 (3 MARKS)

The periodic table acts like a list that shows all of the elements that currently exist. Consider the following elements:

Ca Vanadium Phosphorus Si Be

a Which elements would be found in the same period? (1 MARK)

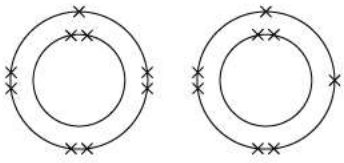
b Which elements would be found in the same group? (1 MARK)

c Which element would most likely have the greatest atomic radius? (1 MARK)

Question 15 (5 MARKS)

Scientists use experimental results to determine properties and characteristics of different elements and molecules.

Consider the following atoms.



	Element A	Element B
Ionisation energy	1313.9 kJ	1681 kJ
Atomic radius	48 pm	42 pm
Period		

A scientist conducting an experiment recorded the above data.

a Complete the table above. (2 MARKS)

b Based on the information presented in the table, determine the name of elements A and B. Explain. (3 MARKS)

Question 16 (5 MARKS)

Sodium and potassium are elements that share similar characteristics. For example, both exhibit a silvery look and interact intensely with water.

a Identify the core charge of each element. (1 MARK)

b Compare the atomic radius of both elements. (2 MARKS)

c Of the two, which element is more likely to have a lower ionisation energy? Explain. (2 MARKS)

Question 17 (5 MARKS)

Fluorine and bromine can both be found as gases in the atmosphere. Fluorine is highly flammable and inhaling bromine can cause severe health problems.

a From which element would it be hardest to remove a valence electron? Explain. (2 MARKS)

b After conducting multiple tests, it was discovered that the elements had an electronegativity score of 3.98 and 2.96, however the scientist was unsure which measurement belonged to which element. Using your understanding, allocate each piece of data to the appropriate element. Explain your reasoning. (2 MARKS)

c Suppose both elements were within proximity of a floating electron that did not belong to any atom. Which of the two elements would be more likely to attract the electron to itself? (1 MARK)

Multiple lessons

Question 18 (3 MARKS)

Aluminium is widely used in everyday life in food packaging materials to building products.

a How many valence electrons are found in aluminium? (1 MARK)

b The ionisation energy of boron is measured as 801 kJ whereas for aluminium it is 578 kJ. Explain why the ionisation energy of boron is slightly higher than that of aluminium. (2 MARKS)

Question 19 (5 MARKS)

Two elements were selected at random, both belonging to period 4. Element A contains 32 electrons and element B contains 20 electrons.

a If both atoms were neutrally charged, identify the number of protons in each element. (1 MARK)

b Write the electron configuration for each element. (2 MARKS)

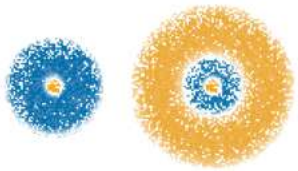
c Of the two elements, which would have a smaller atomic radius? Explain. (2 MARKS)

Key science skills

Question 20 (4 MARKS)

Scientists have developed a microscope that is able to visualise atoms, however not the way we're used to seeing it in textbooks.

To test the machine, the scientist obtained two elements lithium and sodium to be viewed under the microscope. The scientist wanted to see whether the clarity of the image would increase if the atom was left under the microscope for a longer period of time.



a Identify the dependent variable. (1 MARK)

b The results from the test can be seen below:

Test 1

	1 min	5 mins	10 mins
Clarity	Not very clear	Clear	Semi-clear

Test 2

	1 min	5 mins	10 mins
Clarity	Clear	Clear	Not very clear

i What kind of data was collected in the experiment? (1 MARK)

ii Comment on the reproducibility of the results. (2 MARKS)

Exam-style question hints

14a: Periods are viewed as rows in the periodic table. **14b:** Groups are viewed as columns in the periodic table. **14c:** The atomic radius increases down a group. **15a:** Periods reflect energy shells. **15b:** Elements with higher core charges have a greater tendency to attract valence electrons. **16a:** Core charge is calculated based on the number of protons and inner electrons. **16b:** Atomic radius is impacted by the number of protons and energy shells. **16c:** Ionisation energy indicates how difficult it is for a valence electron to be removed from an atom. **17a:** The attraction between the positive nucleus and valence electrons is what keeps the valence electrons in the atom. **17b:** Elements with a strong ability to attract electrons are more electronegative. **17c:** Elements with a higher electronegativity are able to more strongly attract electrons. **18a:** Outer shell electrons are also known as valence electrons. **18b:** Ionisation energy depends on the atomic radius and core charge. **19a:** Neutral atoms have no overall charge. **19b:** Electron configurations show the electron arrangement of elements. **19c:** Atomic size decreases across a period. **20a:** The independent variable is changed by the experimenter. **20bi:** Data can be categorised based on the information collected. **20bii:** Reproducibility relates to the consistency of experimental data between experiments.

2D THE PERIODIC TABLE - PART 2

In this lesson, we will further investigate how the periodic table is organised in relation to reactivity, metallic character and blocks.

2A The atom | 2B Defining elements | 2C The periodic table - part 1 | 2D The periodic table - part 2 | 2E Relative mass | 2F The mole - part 1 | 2G The mole - part 2 | 2H Chemical formulas and equations

Study design dot point

- the periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic/non-metallic character and reactivity) of elements

Key knowledge units

Reactivity	1.1.4.4
Metallic character	1.1.4.5
Blocks on the periodic table	1.1.4.6

Key terms and definitions

Reactivity tendency of atoms to undergo a chemical reaction

Octet rule the tendency of atoms to have eight electrons in their valence shell

Noble gases unreactive gases found in group 18 of the periodic table

Metallic character chemical properties associated with metals

Metalloids elements which possess properties that are in between those of metals and non-metals

Subshells subdivision of electron shells into similar energy levels

Blocks assortment of elements into discrete categories based on which subshell their valence electrons are found in

Lesson links

This lesson builds on:

- 2C The periodic table - part 1

The periodic table also categorises elements based on their metallic and reactive properties.

Reactivity 1.1.4.4

OVERVIEW

The reactivity of an element depends on various factors.

THEORY DETAILS

In the previous lesson, we learned about the properties affecting an atom's ability to attract or lose electrons including the number of electron shells present, the number of valence electrons in the outer shell, the size of the atomic radius and the core charge of the atom. These concepts are important as it helps us to understand the **reactivity** of different elements. Reactivity describes how easy it is for an atom of an element to gain or lose electrons. The easier it is for this to occur, the more reactive the element.

The main reason why atoms gain or lose electrons is to become more stable. Atoms are most stable when they have an electron arrangement that allows them to satisfy the **octet rule**. According to this rule, atoms are most stable when they have eight valence electrons. Therefore, atoms will gain or lose electrons in order for this to occur.

Whether atoms gain or lose electrons depends on the number of valence electrons already present in the atom. For example, metallic elements found in group 1 and 2 which contain one and two valence electrons respectively, would rather lose their valence electrons to have a full valence shell as shown in figure 1. In contrast, non-metallic elements in groups 16 and 17 which contain six and seven valence electrons respectively, would prefer to gain one or two more electrons to satisfy the octet rule as shown in figure 1.

Based on the previous lesson we know that atomic radius, electron shielding and core charge can all impact on how easy it is for atoms to lose or gain electrons.

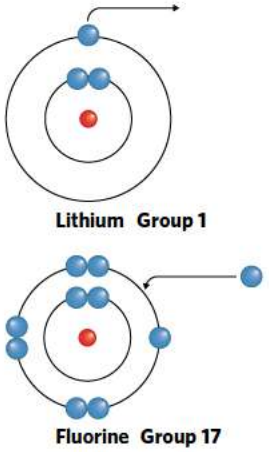


Figure 1 Behaviours of atoms in different groups of the periodic table.

We know that when we move down a group, it becomes easier to lose valence electrons. This can be seen also through the changes in ionisation energy, where elements further down a group have lower ionisation energy than those at the top of the group. As such, as we move down a group, elements become more reactive as it is easier for atoms to lose valence electrons.

Reactivity across a period is interesting, as the criterion changes as we move across different types of elements. While we define reactivity of metals as how easily metals lose electrons, we define reactivity of non-metals by how easily non-metals gain electrons. As such, groups 16 and 17 elements can also be described to be highly reactive, as they have a great tendency to gain electrons based on their high electronegativity and core charge. A summary of the factors affecting reactivity can be seen in table 1.

Table 1 Summary of the impact of different properties on reactivity.

Property	Effect on reactivity	Relative Strength
Increase in atomic radius	↑ for elements that tend to lose electrons ↓ for elements that tend to gain electrons	A large atomic radius allows valence electrons to be easily lost from the valence shell however reduces the ability of atoms to attract electrons due to the shielding electrons.
Increase in core charge	↓ for elements that tend to lose electrons ↑ for elements that tend to gain electrons	An increase in core charge allows atoms to better hold onto their valence electrons and gain electrons more easily.
Increase in electronegativity	↓ for elements that tend to lose electrons ↑ for elements that tend to gain electrons	Electronegativity relates to an atom's ability to attract electrons. Atoms with high electronegativity are more likely to gain electrons.

There is one special group on the periodic table that includes the least reactive elements. These elements are found in group 18 and are also referred to as the **noble gases**.

Metallic character 1.1.4.5

OVERVIEW

Some metallic elements can easily lose valence electrons.

THEORY DETAILS

Elements in group 1 and 2, also known as the alkali and alkaline earth metals respectively, exhibit similar physical properties. For example, both types of metals are shiny, have a silvery appearance and are good conductors of electricity. Scientists define a metal in the periodic table as an element that readily loses electrons to form a positively charged ion. Elements that have a similar physical appearance and chemical properties to the elements in group 1 and 2 possess high **metallic character**. Metals will be further investigated in the next chapter.

To identify the trend in metallic character, the elements can first be organised by how easily electrons are lost. As we've learned earlier, it is easier to remove electrons from an atom as we move down a group, than when we move across a period from left to right. As a result, metallic character increases as we move down a group and decreases as we move across a period from left to right. This can be summarised in figure 2.

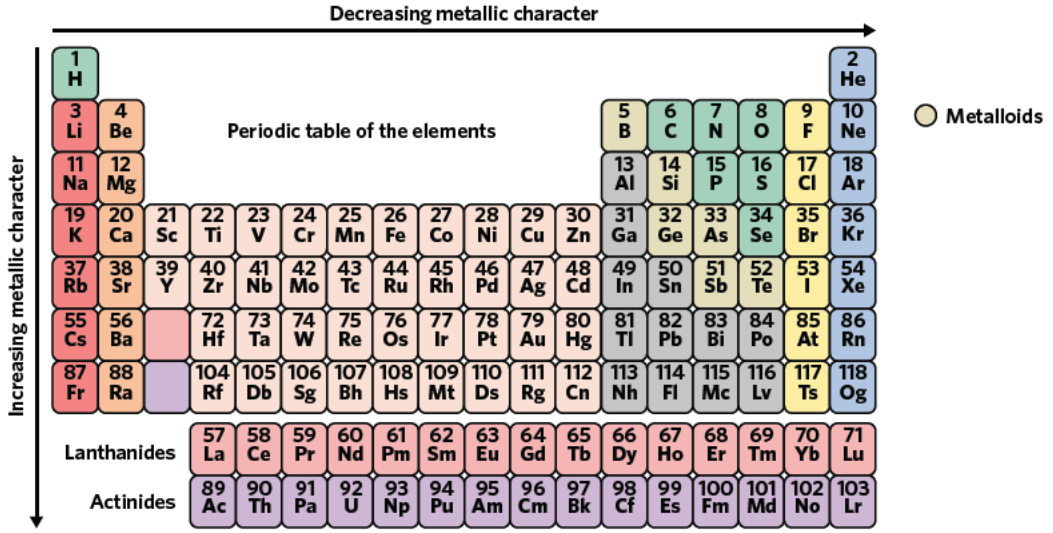


Figure 2 Trend in the metallic character of elements in the periodic table.

Metalloids are an interesting group of elements that display both a mixture of metallic and non-metallic properties. Silicon, considered a metalloid, is able to conduct electricity however is not shiny or silvery in appearance. Some other metalloids include boron, arsenic and tellurium as highlighted in figure 2.

Blocks on the periodic table 1.1.4.6

OVERVIEW

Elements can be grouped based on their highest energy subshell.

THEORY DETAILS

As we learned in lesson 2A, the electron configuration of atoms involves a complex system that involves electrons filling **subshells** of lower energy to higher energy. Combined with our understanding of the information reflected by the periods of a periodic table, we can now appreciate the organisation of elements into blocks as shown in figure 3.

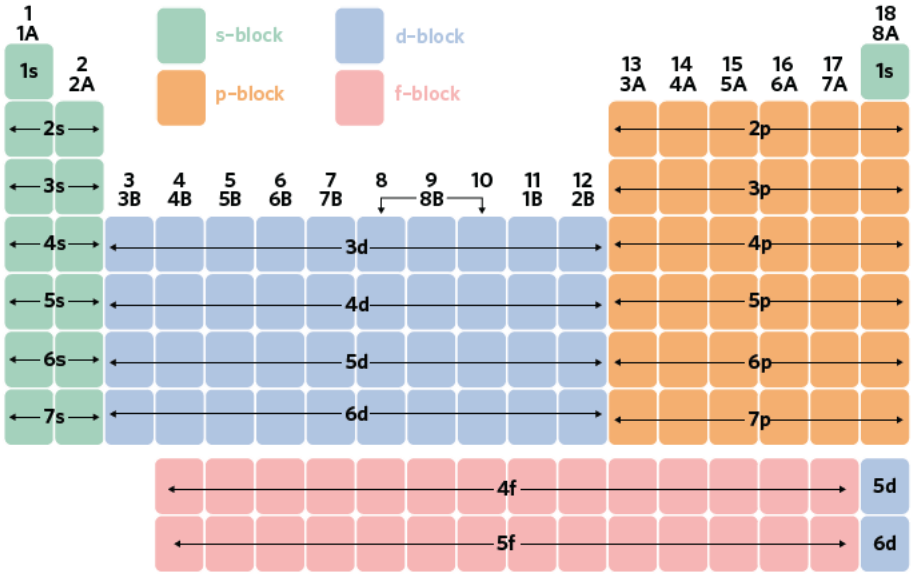


Figure 3 Elemental blocks in the periodic table.

The periodic table is also organised into different categories called **blocks**. Specifically, elements are placed into blocks according to which subshell their highest energy, valence electrons fill. As we can see in figure 3, there are four different blocks, each corresponding to a subshell: s,p,d or f. Elements that fall within the s-block will have the s subshell as its most outer subshell. For example, lithium, which is found in period two, falls within the s-block and has a 2s subshell as its highest energy subshell. This pattern is the same for all of the other blocks.

Worked example 1

In which block would carbon be found on the periodic table? Include the electron configuration of carbon to support your answer.

What information is presented in the question?

The element in question is carbon.

What is the question asking us to do?

Find the block in which carbon is found in the periodic table.

What strategy(ies) do we need in order to answer the question?

1. Write the electron configuration for carbon.
2. Identify the last subshell in carbon as shown in the electron configuration.

Answer

As a neutral atom, carbon contains a total of six electrons. Therefore the electron configuration for carbon is $1s^2 2s^2 2p^2$. As a result, carbon would be found in the p-block.

Theory summary

- Reactivity reflects the ease at which an atom will lose or gain electrons to satisfy the octet rule.
- The octet rule outlines that elements require a total of eight electrons in their valence shell to become stable.
- Metals have a tendency to lose electrons whereas non-metals have a tendency to gain electrons.
- Metallic character decreases as we move across a period.
- Metalloids have characteristics of both metals and nonmetals.
- There are four blocks on the periodic table: s,p,d and f.
- Blocks on the periodic table reflect the highest energy subshell found in the atoms of an element.

2D QUESTIONS

Theory review questions

Question 1

Atoms that can easily lose electrons are highly reactive.

A True

B False

Question 2

Atoms that easily gain electrons are highly reactive.

A True

B False

Question 3

Reactivity of metals _____ from group 1 to group 3.

A increases

B decreases

Question 4

Metals

A contain five or more valence electrons.

B are able to easily lose electrons.

Question 5

Metallic character increases as we move across a period.

A True

B False

Question 6

Metalloids are the most metallic elements on the periodic table.

A True

B False

Question 7

Elements in the s block

A have the same number of valence electrons.

B have the same outer subshell.

Question 8

Elements in the same period belong in the same block.

A True

B False

Question 9

An element found in the p block

A will have the same number of energy shells.

B will have a maximum of six electrons occupying their outer subshell.

Question 10

Oxygen can be found in block

A s.

B p.

Deconstructed questions

Use the following information to answer questions 11–13.

Barium and magnesium both exhibit similar chemical properties. For example, barium and magnesium both react with water.

Question 11 (1 MARK)

Which of the following statements about these elements is true?

A Barium and magnesium are found in period 1.

B Barium is found in period 6 and magnesium is found in period 3.

C Barium is found in group 6 and magnesium is found in group 3.

D Barium and magnesium are found in group 2.

Question 12 (1 MARK)

In a reaction, barium would be most likely to

A gain electrons.

B lose electrons.

C gain protons.

D lose protons.

Question 13 (3 MARKS)

For certain elements, a way in which we test their reactivity is to allow them to come into contact with water and make observations about the interaction that occurs. If a reaction were to occur, there may be the formation of bubbles and/or sparks, which would be observed relatively quickly if the element was highly reactive.

Both magnesium and barium react with water in this way. Of the two, which element would show signs of a reaction the quickest? Explain with reference to the periodic table.

Exam-style questions

Within lesson

Question 14 (2 MARKS)

Emilia was given the following samples and told to conduct an experiment to determine the predominant element found in each.

Based on the data obtained, she was able to narrow it down to the following elements: Mg and P.

Which of the samples would most likely contain the element Mg? Explain.



Sample A **Sample B**

Image: Fokin Oleg/Shutterstock.com

Question 15 (3 MARKS)

In a game of 'chemical elements' celebrity heads, a student asked a series of questions to determine the name of their given element. Below is a series of questions and answers given during the game:

Student: Is the element highly reactive?

Class: Yes

Student: Is the element in block p?

Class: No

Student: Is the element found in periods 1–3?

Class: No

The class gave the student a final clue, saying that it could be one of the following:

Fr C Na Ne

Which element could it be? Explain.

Question 16 (4 MARKS)

Consider three elements, where the last subshell of each electron configuration is 3s, 4s and 6s.

a Suppose each subshell had a single electron. Which element would be the most reactive? Explain. (2 MARKS)

b Describe the metallic character of all three elements. (2 MARKS)

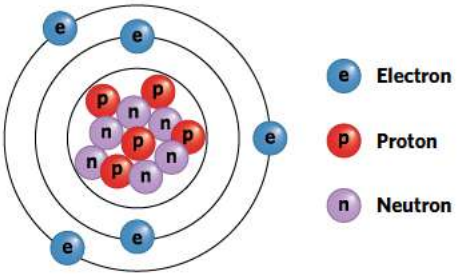
Multiple lessons

Question 17 (2 MARKS)

Consider the following element.

a Write the electron configuration for this element. (1 MARK)

b In which block would we find this element on the periodic table? (1 MARK)



Question 18 (7 MARKS)

The atomic radius describes the distance between the nucleus and valence shell of an atom.

a Describe how the number of shells affects the atomic radius of an atom. (1 MARK)

b Consider periods 2 and 4 of the periodic table.

i Which period would you expect to have overall higher atomic radii? (1 MARK)

ii Describe the trend in atomic radius moving across a period. (2 MARKS)

c Bromine and chlorine are both found in Group 17 on the periodic table. Upon investigation, it was found that the atomic radius of each element is:

	Chlorine	Bromine
Atomic radius (picometres)	102	120

Given this information, identify the atom that would be most reactive. Explain. (3 MARKS)

Question 19 (8 MARKS)

Two students are preparing for a presentation to their chemistry class. The notes the students use for the presentation are composed of a series of statements and justifications in order to tackle the misconceptions of the class. A segment of their notes is given.

Complete the table by identifying whether each statement is correct or incorrect and give a justification for your choice.

Statement	Correct/Incorrect	Justification
Atoms with high reactivity are expected to have high electronegativity.		
Metals are more reactive than non-metals.		
Metals can be found on the left side of the periodic table.		
First ionisation energy is highest for the noble gases.		

Key science skills

Question 20 (9 MARKS)

A scientist discovers a beaker hidden in a cupboard containing a solution of aqueous sodium chloride (NaCl). In order to further explore the properties of the sodium chloride, the scientist designs a practical investigation.

a The first step of this investigation is researching the substance.

i The scientist must classify the substance as belonging to one of the following categories: atom, element, compound mixture. Which category is most appropriate? (1 MARK)

ii The research conducted requires the period and group to be identified for each component of the substance. Identify both the period and group of Na and Cl. (2 MARKS)

Now the scientist must begin executing an experiment. A range of elements in the same period as sodium and chlorine are analysed and the following results are obtained.

Sample	A	B	C	D	E
First ionisation energy	494 kJ mol^{-1}	577 kJ mol^{-1}	785 kJ mol^{-1}	997 kJ mol^{-1}	1250 kJ mol^{-1}

b Explain the trend shown in the table, assuming the samples are listed in order of ascending mass number, with reference to core charge. (3 MARKS)

c Identify the dependent variable in this experiment. (1 MARK)

d A fellow colleague of the scientist notes that the scales used to weigh the sodium chloride prior to conducting the first ionisation energy analysis were not calibrated correctly. What type of error would this induce? Explain your answer. (2 MARKS)

Exam-style question hints

14: Metals have a distinct physical appearance. **15:** The atomic structure of an element can affect its properties. **16a:** Elements in the s block belong to the first two groups in the periodic table. **16b:** Metallic character is synonymous with acts like a metal. **17a:** Electrons fill an atom in a particular order. **17b:** Blocks on the periodic table are based on the outermost subshell. **18a:** Electron shells increase the size of an atom. **18b:** The period number indicates the number of electron shells. **18bii:** The number of protons in the nucleus increases across a period. **18c:** Each element behaves differently during a chemical reaction. **19:** The atomic structure of an element impacts the properties of the element. **20ai:** Substances are classified based on their composition. **20aii:** Periods and groups are like coordinates for an element. **20b:** The ability to remove electrons depends on the strength of attraction to the nucleus. **20c:** The dependent variable is the variable affected by the change in the independent variable. **20d:** Errors are categorized differently depending on their effect on the experiment.

2E RELATIVE MASSES

In this lesson, we will learn about how the relative masses of elements are calculated.

2A The atom	2B Defining elements	2C The periodic table - part 1	2D The periodic table - part 2	2E Relative masses	2F The mole - part 1	2G The mole - part 2	2H Chemical formulas and equation

Study design dot points

- the relative isotopic masses of elements and their representation on the relative mass scale using the carbon-12 isotope as the standard; reason for the selection of carbon-12 as the standard
- determination of the relative atomic mass of an element using mass spectrometry (details of instrument not required)

Key knowledge units

Relative masses of elements	1.1.13.1
Mass spectrometry	1.1.14.1
Relative molecular and formula mass	1.1.13.2

Key terms and definitions

Atomic mass unit (amu) unit of mass used to express atomic masses, relative to $\frac{1}{12}$ of the mass of a single carbon-12 atom

Relative isotopic mass mass of a single atom of an isotope of an element relative to $\frac{1}{12}$ of the mass of a single carbon-12 atom

Relative abundance percentage of a particular isotope found in a naturally occurring sample of an element

Relative atomic mass (A_r) weighted mean of the relative isotopic masses of all isotopes of an element

Mass spectrometry analytical technique used to measure the mass of ions relative to their charge

Mass-to-charge ratio the mass of an ion divided by its charge

Relative molecular mass (M_r) average mass of one molecule of a substance relative to $\frac{1}{12}$ of the mass of a single carbon-12 atom

Molecular formula chemical formula which gives the type and number of atoms of each element in a molecule

Relative formula mass mass of a unit of a substance, particularly an empirical unit, relative to $\frac{1}{12}$ of the mass of a single carbon-12 atom

Lesson links

This lesson builds on:

- 2B Defining elements

The relative masses of elements shown on the periodic table takes into account existing isotopes.

Relative masses of elements 1.1.13.1

OVERVIEW

Even though atoms are invisible to the naked eye, they still have a 'mass'.

THEORY DETAILS

As atoms are so small, it is difficult and inconvenient for chemists to measure them in kilograms, or other SI units. Consequently, chemists use the **atomic mass unit (amu)** to measure the mass of atoms. The atomic mass unit is a relative measurement which means that the mass of every atom is compared with the mass of a single atom of the carbon-12 isotope, ^{12}C, which is assigned a value of 12 amu. The ^{12}C isotope has six protons and six neutrons in the nucleus, with the mass of electrons negligible in comparison. As a result, we can say that this ^{12}C isotope has approximately 12 protons and neutrons in the nucleus, and has an atomic mass of 12 amu.



The relative atomic mass of each element on the periodic table appears as a decimal rather than a whole number because each relative atomic mass is found by averaging the masses of all naturally occurring isotopes. This means that the number of protons and neutrons calculated for each element is really just an average across all the different isotopes!

As we learnt in lesson 2B, there are different forms of the same element called isotopes which contain the same number of protons but a different number of neutrons. This means that each isotope of an element has a slightly different mass which can be compared with the mass of a single carbon-12 atom. We refer to this property as the **relative isotopic mass**. Each isotope of an element takes up a different percentage of the overall mass of an element which is referred to as the **relative abundance** of each isotope. A rarer isotope with a lower relative abundance contributes less to the overall mass of an element than a more common isotope with a higher relative abundance. Another term for the overall mass of an element is the **relative atomic mass** which is the weighted average of all the individual relative isotopic masses. On the periodic table, the relative atomic mass of each element is displayed underneath the element symbol and name as shown in figure 1.

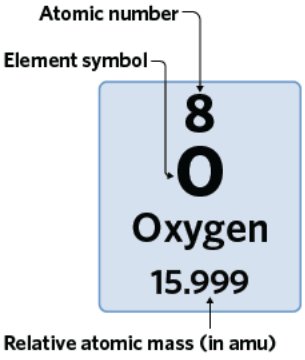


Figure 1 The representation of an element on the periodic table.

As the relative atomic mass (A_r) is the weighted mean of the relative isotopic masses of an element, we can calculate the relative atomic mass of an element using the formula:

$$A_r = \frac{(\textit{relative abundance (as a percentage)} \times \textit{relative isotopic mass}) + (\textit{relative abundance (as a percentage)} \times \textit{relative mass}) + \ldots}{100}$$

Worked example 1

It is known that the element chlorine has two stable isotopes. Chlorine-35 has a relative abundance of 76% and chlorine-37 has a relative abundance of 24%. Calculate the relative atomic mass of chlorine.

What information is presented in the question?

We know that there are two isotopes of chlorine with relative isotopic masses of 35 and 37 which have a relative abundance of 76% and 24% respectively.

What is the question asking us to do?

Calculate the relative atomic mass.

What strategy(ies) do we need in order to answer the question?

1. Identify the formula for relative atomic mass.
2. Substitute the values given from the question into the formula.
3. Calculate the relative atomic mass.

Answer

Isotope	Relative abundance (%)
35	76
37	24

The formula for relative atomic mass is:

$$A_r = \frac{(\textit{relative abundance} \times \textit{relative isotopic mass}) + (\textit{relative abundance} \times \textit{relative isotopic mass})}{100}$$

We can substitute the values into the equation.

$$A_r = \frac{(76 \times 35) + (24 \times 37)}{100}$$

$$= 35.5$$

Calculating the relative abundance of isotopes from the relative atomic mass

Sometimes, we may not be given the relative abundance of each isotope however if we are given the relative atomic mass of the element we will be able to calculate the relative abundance.

Consider the element bromine, which has a relative atomic mass of 79.9. It exists as two isotopes, ^{79}Br and ^{81}Br. Although we don't know what the relative abundance of either bromine isotope is, we know that the total relative abundance of isotopes must always add up to 100%. As such, the relative abundance of ^{79}Br can be denoted as x and the relative abundance of ^{81}Br can be denoted as $100 - x$. This is shown in table 1.

If we substitute the information in table 1 and the relative atomic mass for bromine (as shown on the periodic table) into the A_r formula, we can see that:

$$79.9 = \frac{(x \times 79) + ((100 - x) \times 81)}{100}$$

$$79.9 = \frac{79x + 8100 - 81x}{100}$$

$$79.9 \times 100 = \frac{79x + 8100 - 81x}{100} \times 100$$

$$7990 = 79x + 8100 - 81x$$

$$7990 = -2x + 8100$$

$$7990 - 8100 = -2x + 8100 - 8100$$



Don't confuse relative abundance with relative isotopic mass! Sometimes, the lightest isotope is the most common and the heaviest isotope is the least common.

Table 1 Calculating the relative abundance of bromine isotopes.

Isotope	Relative abundance (%)
^{79}Br	x
^{81}Br	$100 - x$
total	100

$-110 = -2x$

$\frac{-110}{-2} = \frac{-2x}{-2}$

$55 = x$

If we then substitute the value obtained for x back into table 1 below, we can determine that the relative abundance of ^{79}Br is 55% and the relative abundance of ^{81}Br is 45%.

Table 2 Calculating the relative abundance of bromine isotopes.

Isotope	Abundance (%)
^{79}Br	55
^{81}Br	100 − 55 = 45
total	100

Worked example 2

Europium was discovered by a French chemist in the late 1800s. It was shown that Europium has two isotopes, ^{151}Eu and ^{153}Eu. If the relative atomic mass is 151.96, calculate the relative abundance of each isotope.

What information is presented in the question?

We know there are two isotopes of Eu and that the A_r of Eu is 151.96.

What is the question asking us to do?

Calculate the relative abundance of each isotope.

What strategy(ies) do we need in order to answer the question?

1. Identify the formula for relative atomic mass and make x represent the unknown relative abundance.
2. Substitute the values given from the question.
3. Calculate the relative abundance of one isotope.
4. Use the value obtained in step 3 to determine the abundance of the second isotope.

Answer

Isotope	Relative abundance (%)
151	x
153	$100 - x$

The formula for relative atomic mass is:

$$A_r = \frac{(\text{relative abundance} \times \text{relative isotopic mass}) + (\text{relative abundance} \times \text{relative isotopic mass})}{100}$$

We can substitute the values into the equation.

$$151.96 = \frac{(x \times 151) + ((100 - x) \times 153)}{100}$$

$$151.96 = \frac{151x + 15\,300 - 153x}{100}$$

$$151.96 \times 100 = \frac{151x + 15300 - 153x}{100} \times 100$$

$15196 = 151x + 15300 - 153x$

$15196 = -2x + 15300$

$15196 - 15300 = -2x + 15300 - 15300$

$-104 = -2x$

$\frac{-104}{-2} = \frac{-2x}{-2}$

$52 = x$

Hence, the relative abundance of ^{151}Eu is 52.0% and the relative abundance of ^{153}Eu is:

$100 - 52.0 = 48.0\%$

Mass spectrometry 1.1.14.1

OVERVIEW

Chemists use a special technique called mass spectrometry to determine the relative atomic mass of elements.

THEORY DETAIL

As we have seen above, the isotopes of elements exist in different proportions. We can measure the relative abundance of isotopes using an analytical technique known as **mass spectrometry**. In the mass spectrometer, samples are converted into cations which move through the instrument in a particular way depending on the relationship between the mass of the sample and its charge, also known as the **mass-to-charge ratio**. The mass to charge ratio is represented as m/z. The results of a mass spectrometry experiment are visualised on a mass spectrum as shown in figure 2.

Calculating the relative atomic mass of elements

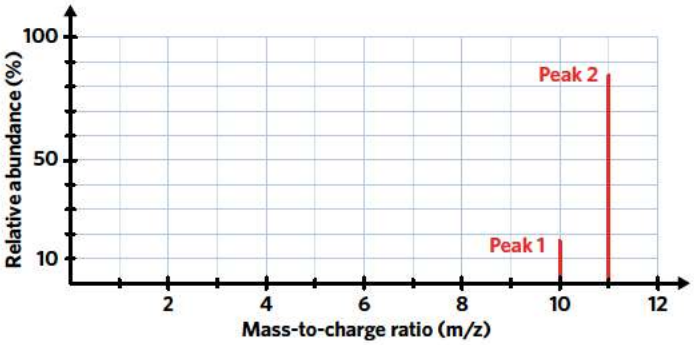


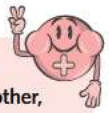
Figure 2 The mass spectrum of boron.

Each 'peak' on the spectrum corresponds to an isotope of boron. The height of a peak indicates how abundant the isotope is, as a proportion of the total mass of the sample. The x-axis indicates the mass-to-charge ratio, where the heavier the atom, the greater the mass to charge ratio. Using this information, we can calculate the A_r of boron by:

$$A_r = \frac{(18.7 \times 10) + (81.3 \times 11)}{100}$$

$A_r = 10.8$ amu

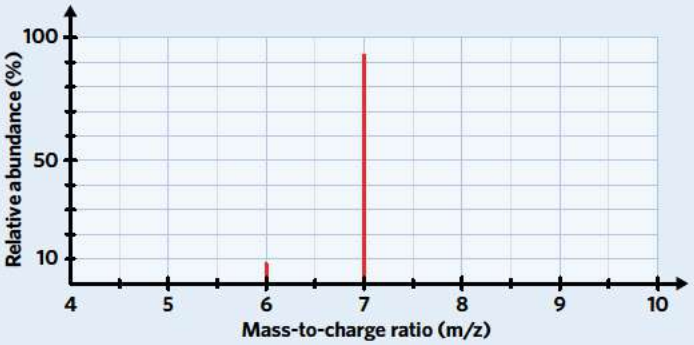
Since boron 11 is the most abundant isotope, we can see that the A_r of boron is much closer to 11 than 10. As such, the atomic masses presented in the periodic table are weighted towards the most abundant isotope of a particular element.



If the A_r is closer to one isotope than another, the isotope closest to the A_r has a higher relative abundance.

Worked example 3

A sample of a single, monoatomic element was analysed using mass spectrometry to produce the spectrum given. Using the results, determine which element was analysed.



Mass to charge ratio	Relative abundance (%)
6.02	7.5
7.02	92.5

What information is presented in the question?

We have a mass spectrum with two peaks and a table of values showing the exact numbers represented by each peak. We also know that the element is monatomic, and only a single element was analysed.

What is the question asking us to do?

Determine the identity of the element.

What strategy(ies) do we need in order to answer the question?

1. Substitute the mass-to-charge ratio and relative abundance of each isotope into the A_r formula to calculate the relative atomic mass of the element.
2. Use the relative atomic mass and the periodic table to determine which element was analysed.

Answer

$$A_r = \frac{(7.5 \times 6.02) + (92.5 \times 7.02)}{100}$$

$A_r = 6.9$

Based on the calculations, it can be seen that the relative atomic mass of the unknown element is approximately 6.9 amu. According to the periodic table, the element with a relative atomic mass of 6.9 amu is lithium.

Calculating the relative abundance of isotopes from mass spectra

Now that we know the basic calculations involving relative atomic mass, we can use this understanding to determine the relative abundance of different isotopes. Depending on the information given, there are a couple of different ways in which we can do this.

Consider the mass spectrum shown in figure 3. Notice that there is no scale for the y-axis. However, we know that the peaks in the spectrum are relative to each other, and together make up a total of 100%.

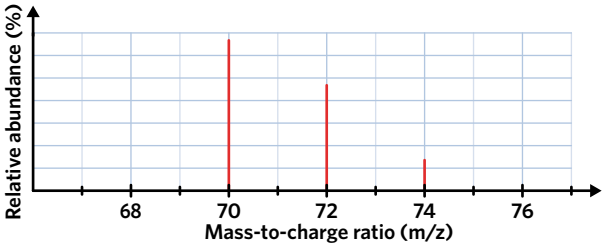


Figure 3 Mass spectrum of an unknown element.

Therefore, what we can do is use a ruler to determine the length of each peak, and use these values to determine the relative abundance of each isotope as seen in figure 4.

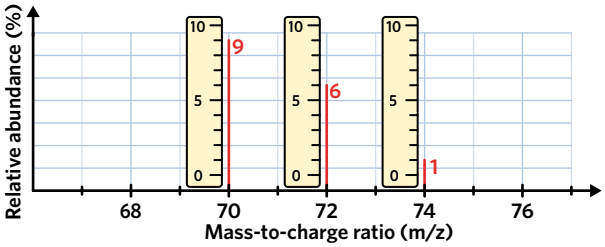
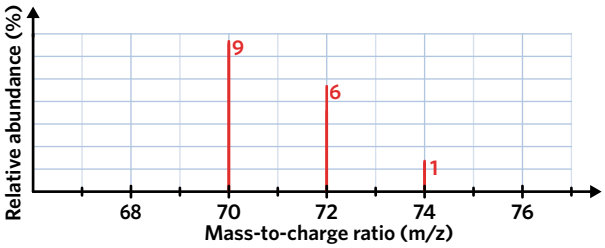


Figure 4 Process of using a ruler to measure the peaks in a mass spectrum.

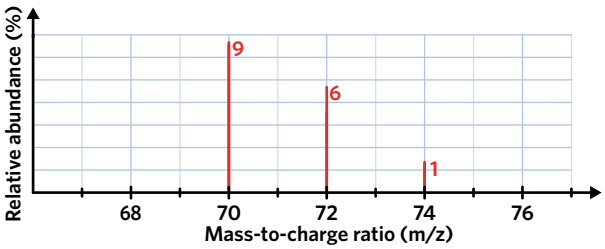
As these peaks are a percentage of the total, we need to first determine the total peak height of all peaks. This would be calculated as shown in figure 5.



$$\textit{Total peak height} = 9 + 6 + 1 = 16$$

Figure 5 Calculation of the total peak height of all peaks in a mass spectrum.

Once we have calculated the total height, we can determine the abundance of each peak as shown in figure 6.



$$\textit{\% abundance of peak 70} = \frac{9}{16} \times 100 = 56.25\ \%$$

$$\textit{\% abundance of peak 72} = \frac{6}{16} \times 100 = 37.50\%$$

$$\textit{\% abundance of peak 74} = \frac{1}{16} \times 100 = 6.25\%$$

Figure 6 Calculation of the percentage abundance of each peak.

Relative molecular and formula mass 1.1.13.2

OVERVIEW

Using the relative atomic mass of elements, we can calculate the relative mass of molecules.

THEORY DETAILS

As we know, elements are able to bond together to form molecules. We can use the relative atomic mass of single elements to calculate the **relative molecular mass** of molecules, represented as M_r. To calculate this, we need to first understand the elements and number of atoms of each element present in a molecule which is shown in the **molecular formula**.

For example, the molecular formula for water is H_2O. We can see that there are two elements present: hydrogen and oxygen. The subscript after the chemical symbol H indicates that there are two hydrogen atoms present, and only one oxygen atom present in every water molecule. As such, in order to calculate the relative molecular mass for water, we would sum the atomic mass of all atoms present in the molecule as shown in figure 7.

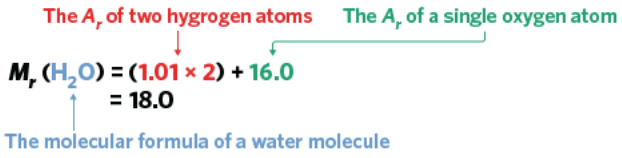


Figure 7 Calculation of the relative molecular mass of water.

In Chemistry we have a special way of communicating ideas, one of which involves how we express calculations. As we can see in the example above, we use the M_r symbol to show what it is that we are looking for, and within the bracket indicate the substance we are trying to find out about, both of which are found on the left hand side of the equation. On the right hand side, we write out the numerical calculations leading to a solution for the relative molecular mass. Brackets are also used to separate symbols or numbers from one another. Figure 8 shows the breakdown of this concept.

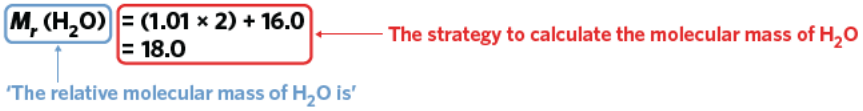


Figure 8 Symbols and brackets used when calculating relative molecular masses.

This concept will apply to all future calculations.

Worked example 4

Hydrogen peroxide, H_2O_2, is used as a hair dye. What is the relative molecular mass of hydrogen peroxide?

What information is presented in the question?

The molecular formula of hydrogen peroxide.

What is the question asking us to do?

Calculate the relative molecular mass of hydrogen peroxide.

What strategy(ies) do we need in order to answer the question?

1. Determine the relative atomic mass of each element present in the molecule.
2. Sum the relative atomic masses of all atoms present in the molecule.

Answer

Element	Number of atoms of the element	Relative atomic mass
H	2	1.0
O	2	16.0

$M_r(H_2O_2) = (2 \times 1.0) + (2 \times 16.0) = 34.0$

Therefore, the relative molecular mass of hydrogen peroxide is 34.0.

As we will learn later, sometimes substances don't exist as molecules, but rather a big continuous lattice structure. In this case, it is inaccurate to use the term molecular mass. As such, we use the term **relative formula mass** to illustrate that we are finding the relative mass based on the formula of the compound. However, it is calculated in the same way as the relative molecular mass.

Theory summary

- The atomic mass unit (amu) compares the mass of every atom, molecule and compound with $\frac{1}{12}$ of the mass of a single carbon-12 atom.
- The relative isotopic mass refers to the mass of a single atom of a particular isotope compared with $\frac{1}{12}$ of the mass of a single carbon-12 atom.
- The relative abundance of an isotope refers to the percentage of that isotope in a naturally occurring sample of an element.
- The mass of each element on the periodic table is referred to as the relative atomic mass.
- The relative atomic mass is calculated by taking into account the relative abundance and relative isotopic masses of each isotope of an element.
- The relative abundance of an isotope can be calculated if the relative atomic mass and relative isotopic masses of an element are known.
- Mass spectrometry is used to determine the relative abundance of different isotopes of an element.
- The relative abundance of isotopes can be calculated using mass spectra by measuring and comparing the lengths of the peaks present.
- A molecular formula gives information about all the elements and the number of atoms of each in one molecule of a substance.
- The relative molecular mass is calculated by adding together all of the relative atomic masses for each atom present in a single molecule.
- The relative formula mass is calculated by adding together all of the relative atomic masses of each atom in the formula of a compound.

2E QUESTIONS

Theory review questions

Question 1

The unit amu represents the

A number of neutrons in an element.

B the relative atomic mass of an element.

Question 2

The relative atomic mass of an element

A depends on the abundance of neutrons in an element.

B depends on the abundance of each isotope of an element.

Question 3

Element Z exists in two forms. ^{14}Z has a relative abundance of 78% and ^{15}Z has a relative abundance of 22%. The relative atomic mass of element Z will be closest to

A 14.

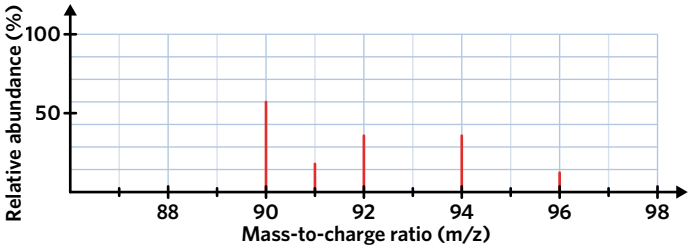
B 15.

Question 4

Zirconium, Zr, is a transitional metal found in the s block of the periodic table. Based on the mass spectrum, Zr has

A 96 isotopes.

B 5 isotopes.



Question 5

According to the mass spectrum in question 4, the least abundant isotope is

A ^{90}Zr.

B ^{96}Zr.

Question 6

Based on the spectrum shown in question 4, the relative atomic mass of Zr would be closest to

A 90.

B 95.

Question 7

Carbon dioxide CO_2 and oxygen O_2 are both molecules. Therefore, the relative molecular mass of both of these molecules

A is the same.

B is different.

Question 8

The relative molecular mass of a molecule is

A the sum of the total number of protons and neutrons found in the molecule.

B the sum of the relative atomic masses of each atom in the molecule.

Question 9

The symbol representing relative molecular mass is

A M_r.

B *MM*.

Question 10

The relative molecular mass of Cl_2 would be

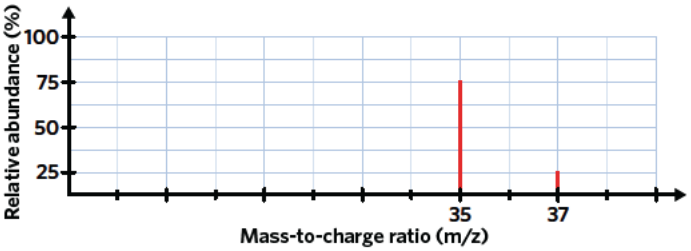
A 35.5.

B 70.1.

Deconstructed questions

The following statement is to be used for questions 11–13.

Mass spectrometry is used to determine the mass of isotopes.



Question 11 (1 MARK)

Based on the mass spectrum, the element in question exists in how many forms?

A 35

B 37

C 1

D 2

Question 12 (1 MARK)

Which of the following formulas would be used to calculate the relative atomic mass of the element?

A $\frac{(35 \times 25) + (37 \times 75)}{100}$

B $\frac{(35 \times 37) + (25 \times 75)}{100}$

C $\frac{(35 \times 75) + (37 \times 25)}{100}$

D $\frac{(35 \times 27) + (37 \times 75)}{100}$

Question 13 (2 MARKS)

What is the name of the element represented by the mass spectrum?

Exam-style questions

Within lesson

Question 14 (3 MARKS)

An element contains three different isotopes with the following relative abundances:

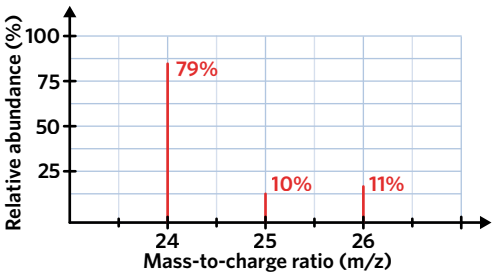
Relative isotopic mass	Relative abundance (%)
19.99	90.92
20.99	
21.99	8.82

a What is the relative abundance of the isotope with a relative isotopic mass of 20.99? (1 MARK)

b Calculate the relative atomic mass of the element. (2 MARKS)

Question 15 (3 MARKS)

An unknown sample was placed into a mass spectrometer and the mass spectrum showed a total of three peaks.

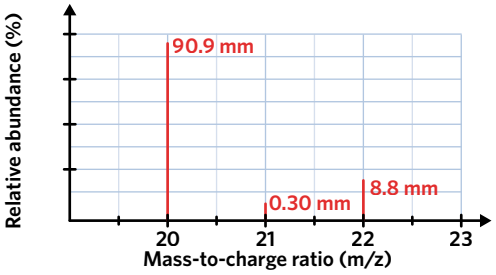


a Calculate the relative atomic mass of the element. (2 MARKS)

b What is the name of the element in the sample? (1 MARK)

Question 16 (5 MARKS)

Students from a university lab came across a sample containing an unknown element. To determine the nature of the unknown element, the students performed mass spectrometry and obtained the mass spectrum provided.



Unfortunately, the program responsible for the development of the mass spectrum was not set to include the percentage abundance of each isotope. Instead, students measured the length of the peaks using a ruler and included this as part of the spectrum.

a Identify the number of isotopes for the unknown element. (1 MARK)

b Calculate the relative abundance of each isotope. (3 MARKS)

c Identify the name of the unknown element. (1 MARK)

Question 17 (8 MARKS)

Lutetium is used in radiation to treat cancer and is known to have two isotopes, ^{175}Lu and ^{176}Lu. When researching Lu, students were unable to find the relative abundance of each isotope however they did find that Lu has an A_r of 174.97 according to the periodic table.

a Based on the relative atomic mass, which isotope would most likely be the most abundant? (1 MARK)

b Oddly, they found that the isotopic mass of ^{175}Lu and ^{176}Lu was 174.94 and 175.94 respectively. Calculate the relative abundance of each isotope. (3 MARKS)

c Draw a mass spectrum for Lu. (3 MARKS)

d Lu can form bonds with Ge to form $LuGe_3$. What is the relative molecular mass of this compound? (1 MARK)

Multiple lessons

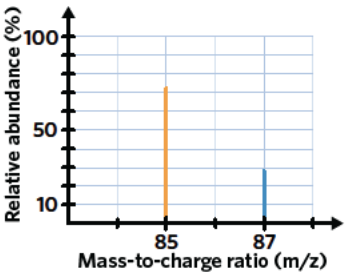
Question 18 (4 MARKS)

Rubidium is an extremely reactive element and therefore is required to be stored in very strict conditions.

a In which block of the periodic table is rubidium found? (1 MARK)

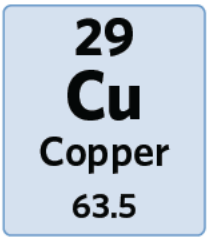
b What is the electron configuration of rubidium's outermost subshell? (1 MARK)

c The mass spectrum of rubidium is shown. Calculate the relative atomic mass of rubidium. (2 MARKS)



Question 19 (6 MARKS)

Consider the following element:



a How many protons would Cu have in its nucleus? (1 MARK)

b Cu is known to have 35 neutrons in its nucleus. What would be the mass number for Cu? (1 MARK)

c Cu is known to exist as two isotopes: ^{63}Cu and ^{65}Cu. Calculate the relative abundance of each isotope. (3 MARKS)

d Copper is able to bond with oxygen to form CuO. What is the relative formula mass? (1 MARK)

Key science skills

Question 20 (5 MARKS)

Rafa was on the verge of discovering a new element. To determine the nature of the element, he conducted multiple rounds of mass spectrometry on his sample. The results were as follows:

Test 1

Characteristic	Data
Number of peaks	2
m/z ratio	71 and 73
Relative abundance of each peak	24.5% for peak 71 and 75.5% for peak 73

Test 2

Characteristic	Data
Number of peaks	2
m/z ratio	71 and 73
Relative abundance of each peak	73.2% for peak 71 and 27.8% for peak 73

Test 3

Characteristic	Data
Number of peaks	2
m/z ratio	71 and 73
Relative abundance of each peak	24.9% for peak 71 and 75.1% for peak 73

a If Rafa stopped the experiment after test 2, comment on the repeatability of the experiment. (2 MARKS)

b Rafa used the data to determine the relative abundance of each peak.

i How many significant figures should be expressed in his answer? (1 MARK)

ii Calculate the relative abundance for each peak. (2 MARKS)

Exam-style question hints

14a: Relative abundance of all isotopes of an element equals 100%. **14b:** Relative atomic mass takes into account the relative isotopic mass of each isotope and its relative abundance. **15a:** Mass/charge ratio of each peak represents the mass of the isotope. **15b:** A/r of an element is closest to the most abundant isotope. **16a:** A mass spectrum shows information about each isotope of an element. **16b:** Peak heights can be used as a point of comparison. **16c:** Each element has its own relative atomic mass. **17a:** Relative atomic mass of an element is skewed towards the most abundant isotope. **17b:** Total relative abundance of all isotopes equals 100%. **17c:** The height of each peak in a mass spectrum represents the relative abundance of an isotope. **17d:** Relative molecular mass is the sum of the relative atomic masses. **18a:** The periodic table is arranged in blocks based on the subshell notation. **18b:** Blocks in which elements are found represent their outermost subshell. **18c:** Relative atomic mass takes into account the relative isotopic mass of each isotope and its relative abundance. **19a:** Atomic number of an element indicates the number of protons in the nucleus. **19b:** Mass number depends on the number of nucleons. **19c:** Relative abundance of isotopes can be calculated from the relative atomic mass. **19d:** Relative formula and molecular masses are calculated in the same way. **20a:** Repeatability of experiments requires the same experimental conditions. **20bi:** Number of significant figures depends on the values used in the calculations. **20bii:** Not all pieces of data should be included when calculating values.

2F THE MOLE - PART 1

In this lesson, we will be learning about the mole – an important concept in Chemistry that will form the foundation of future topics.

2A The atom	2B Defining elements	2C The periodic table - part 1	2D The periodic table - part 2	2E Relative masses	2F The mole - part 1	2G The mole - part 2	2H Chemical formulas and equations

Study design dot point

- the mole concept; Avogadro constant; determination of the number of moles of atoms in a sample of known mass; calculation of the molar mass of ionic compounds

Key knowledge unit

The mole	1.1.15.1

Key terms and definitions

Avogadro's number (N_A) the number of atoms in 12 g of ^{12}C, equal to 6.02×10^{23}

Mole (n) unit of measurement defined as 6.02×10^{23} chemical entities

Amount the number of moles of substance

Lesson links

This lesson builds on:

- 2A The atom

Various models of the atom have been proposed throughout history.

- 2E Relative masses

Atomic, molecular and formula masses are always measured relative to ^{12}C.

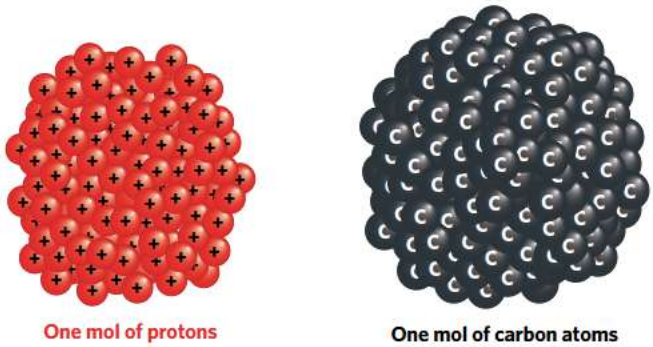
The mole 1.1.15.1

OVERVIEW

It is necessary to express large quantities of minute particles in a simple and convenient manner.

THEORY DETAILS

As mentioned in the previous lesson, carbon-12 (^{12}C) is an important isotope as it acts as a standard for the measurement of relative atomic masses of other elements. Experiments have shown that a 12 g sample of ^{12}C contains a total of $6.02214076 \times 10^{23}$ atoms which is often shortened to 6.02×10^{23}. In the context of our studies, we will use the value 6.02×10^{23}. This value is referred to as **Avogadro's number (N_A)** and is an important numerical constant in chemistry because it defines the value of one **mole**. The mole is denoted by the symbol n and is used to represent the **amount** of any substance whether it be an atom, an ion, a molecule or even a compound. When performing calculations, the mole is written as *mol*. As shown in figure 1, one mol of any substance has the same number of individual entities.



One mol of protons

One mol of carbon atoms

Figure 1 Representation of one mol of protons and one mol of carbon atoms.

As we can see, the mole is essentially a counting unit in the same way that a dozen is used to signify 12 individual entities. The relationship between a mole, Avogadro's number (N_A) and the number of individual entities that is represented by Avogadro's number is shown in figure 2.

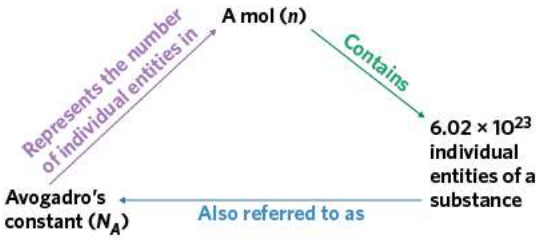
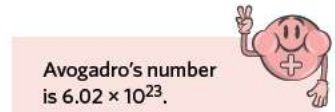
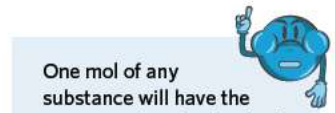


Figure 2 Relationship between a mol, Avogadro's number and 6.02×10^{23}.





As we can see, the total number of particles of substance depends on the amount, in mol, of substance present, as shown in figure 3.

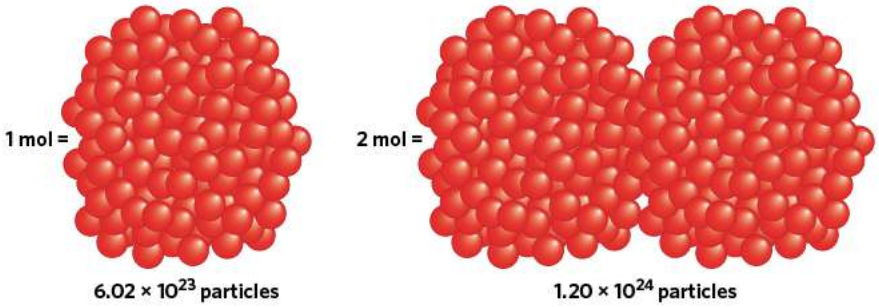


Figure 3 Representation of the number of particles according to the amount, in mol, of substance present.

Increasing the number of mol of substance increases the total number of particles whereas decreasing the number of mol of substance decreases the total number of particles present. We can represent this relationship as shown in figure 4.

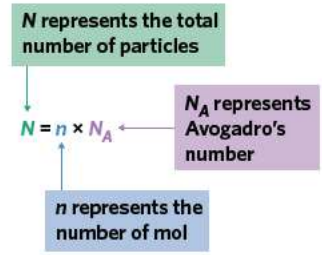


Figure 4 Mathematical equation linking the total number of particles (N), the number of mol (n) and Avogadro's number (N_A).

Worked example 1

How many atoms of copper are present in 1.5 mol of copper?

What information is presented in the question?

The amount, in mol, of copper.

What is the question asking us to do?

Calculate the number of copper atoms present.

What strategy(ies) do we need in order to answer the question?

1. Substitute the known values into the equation.
2. Calculate the number of copper atoms present.

Answer

$N = n \times N_A$

$n = 1.5$ mol

$N_A = 6.02 \times 10^{23}$

$N(\text{Cu}) = 1.5 \times 6.02 \times 10^{23}$

$N(\text{Cu}) = 9.03 \times 10^{23}$

Therefore, there are 9.0×10^{23} copper atoms in 1.5 mol of copper.

Since we know that one mole represents a specific quantity of particles, we can use this information to work backwards to find the number of mole of substance present. As can be seen in figure 5, 6.02×10^{23} particles is equivalent to one mole of substance, whereas 1.20×10^{24} particles is equivalent to two moles of substance.

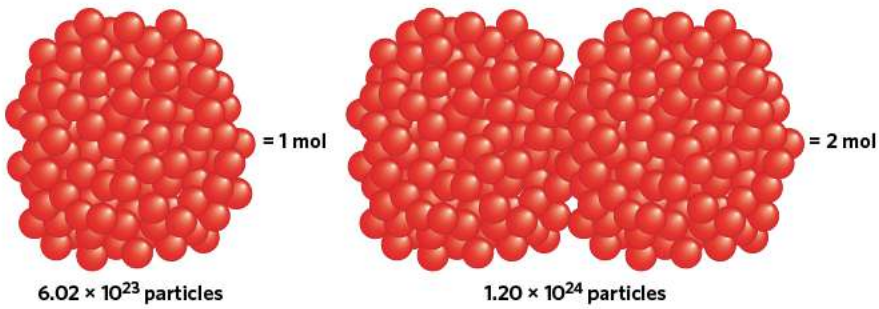


Figure 5 Representation of how the number of particles can be used to find the amount of substance, in mol.

This relationship is shown in figure 6 and is a variation of the equation shown in figure 4.

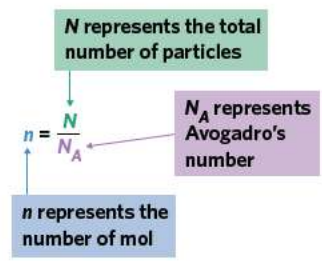


Figure 6 Mathematical equation linking the number of mol (n), the total number of particles (N) and Avogadro's number (N_A).

Worked example 2

In a chemical reaction, scientists determined that 4.5×10^{15} molecules of water (H_2O) were produced. What amount, in mol, of water was produced?

What information is presented in the question?

The number of water molecules.

What is the question asking us to do?

Calculate the amount, in mol, of water produced.

What strategy(ies) do we need in order to answer the question?

1. Substitute the known values into the equation.
2. Calculate the amount of water produced, in mol.

Answer

$$n = \frac{N}{N_A}$$

$N(H_2O) = 4.5 \times 10^{15}$

$N_A = 6.02 \times 10^{23}$

$$n(H_2O) = \frac{4.5 \times 10^{15}}{6.02 \times 10^{23}}$$

$= 7.5 \times 10^{-9}$

Therefore, 7.5×10^{-9} mol of water was produced.

We can use the following concept triangle to help us figure out the type of calculation required to determine the amount of substance, in mol, or the number of particles using Avogadro's number.

To use this, use your finger to cover up the value you are trying to find. Calculate using the remaining variables as indicated by the signs. For example, to find the amount, cover the section of the triangle that says 'amount', which leaves us to perform the calculation $\frac{N}{N_A}$.

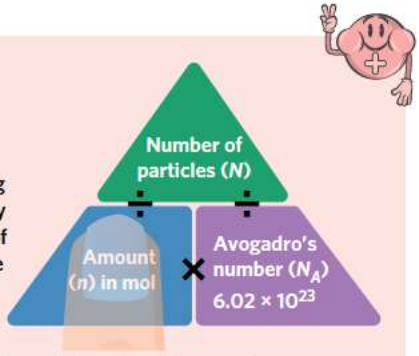


Image: Vasilyeva Larisa/Shutterstock.com

The mole and chemical formulas

Molecular formulas give us information about the elements, and the number of atoms of each, in a molecule. For example, the molecular formula for water, H_2O tells us that for every water molecular, there are two hydrogen atoms and a single oxygen atom. We can also apply the mole concept to molecules.

As shown in table 1, we can use the information given in molecular formulas to determine the number of mole of each element present in a molecule.

Table 1 The relationship between the mole concept and molecular formulas.

Substance	Number of moles of the compound	Number of moles of each element
Water H_2O	1	Hydrogen: 2 mol Oxygen: 1 mol
Oxygen O_2	1	Oxygen: 2 mol
Carbon dioxide CO_2	1	Carbon: 1 mol Oxygen: 2 mol

As we can see, the number of moles depends on what it is that we're referring to, whether whole molecules or elements within molecules.

Theory summary

- The number of atoms found in 12 g of ^{12}C is 6.02×10^{23} which is a numerical constant referred to as Avogadro's number (N_A).
- One mole of any substance is always equal to 6.02×10^{23} particles.
- The mole is expressed in mol.
- The amount of a substance is expressed in mol.
- Since the value of N_A is always known, we can calculate the amount (n) or number of particles (N) using variations of the equation $N = n \times N_A$.

2F QUESTIONS

Theory review questions

Question 1

The mole is based on

A the number of atoms in the world.

B the number of atoms present in 12 g of ^{12}C.

Question 2

A mole is equivalent to

A Avogadro's number.

B the mass number.

Question 3

Avogadro's number is

A 12.

B 6.02×10^{23}.

Question 4

One mole of substance contains

A 12 particles.

B 6.02×10^{23} particles.

Question 5

Half a mole of substance is equivalent to

A two times Avogadro's number.

B half of Avogadro's number.

Question 6

If a mole contains a fixed number of particles, two moles would contain

A half as many particles.

B twice as many particles.

Question 7

One mole of any substance contains the same _____ but could have different _____.

A mass, number of mol

B number of particles, elements

Question 8

One mole of carbon dioxide (CO_2) and one mole of ammonia (NH_3) would contain the same

A number of molecules.

B number of atoms.

Question 9

For every mole of fluorine gas (F_2), there are

A two moles of fluorine atoms.

B half a mole of fluorine atoms.

Question 10

In one mole of hydrogen sulfide (H_2S) and hydrogen fluoride (HF), there would be

A the same number of hydrogen atoms in both substances.

B twice as many hydrogen atoms in hydrogen sulfide as in hydrogen fluoride.

Deconstructed questions

Use the following information to answer questions 11–13.

Cars convert fuel into energy in a process called combustion which releases carbon dioxide (CO_2) as a byproduct. In a reaction to determine the efficiency of a particular fuel type, experimenters measured that 1.22 mol of carbon dioxide was produced for every gram of fuel used.

Question 11 (1 MARK)

The number of particles of carbon dioxide produced in this reaction is dependent on

A the amount of carbon dioxide produced.

B the molar mass of carbon dioxide.

C Avogadro's number.

D the number of carbon atoms present.

Question 12 (1 MARK)

How many molecules of carbon dioxide would be released into the atmosphere if 1.22 mol of carbon dioxide was produced?

A 6.02×10^{23} molecules.

B 7.34×10^{23} molecules.

C 3.67×10^{23} molecules.

D 1.22×10^{23} molecules.

Question 13 (2 MARKS)

How many oxygen atoms were produced as part of this reaction?

Exam style questions

Within lesson

Question 14 (6 MARKS)

Different amounts of substance are required for particular reactions.

a For each of the scenarios, calculate the amount of substance used.

i 4.11×10^{21} molecules of O_2. (1 MARK)

ii 7.05×10^{11} molecules of H_2S. (1 MARK)

iii 1.28×10^{2} atoms of Ar. (1 MARK)

b For each of the scenarios, calculate the:

i number of H_2S molecules in 3.44 mol of H_2S. (1 MARK)

ii number of NO_3 molecules in 0.966 mol of NO_3. (1 MARK)

iii number of F atoms in 1.01 mol of SF_3. (1 MARK)

Question 15 (3 MARKS)

Water can be produced in a chemical reaction involving hydrogen gas and oxygen gas as shown:

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

A student was given two different methods to produce water. One method required the student to use 4.78 mol of hydrogen and the other method required the student to use 5.63×10^{16} molecules of hydrogen gas. Which method would involve the use of more hydrogen gas? Include calculations in your answer.

Question 16 (3 MARKS)

For each of the following comparison, use the symbol > or < to identify the larger quantity.

a 2.11 mol of iron (Fe) atoms __ 3.33×10^{33} molecules of chlorine gas (Cl_2). (1 MARK)

b 5.17×10^{11} molecules of nitrogen dioxide (NO_2) __ 2.98 mol of carbon tetrachloride (CCl_4). (1 MARK)

c 7.01×10^{9} molecules of carbon disulfide (CS_2) __ hydrogen atoms in 2.70×10^{9} molecules of phosphine (PH_3). (1 MARK)

Multiple lessons

Question 17 (3 MARKS)

If left out in the air for too long, some metals can form a layer of rust on the surface. Rust develops based on the reaction between iron and oxygen to produce iron oxide:

$2Fe_{(s)} + O_{2(g)} \rightarrow 2FeO_{(s)}$

a Calculate the formula mass for iron oxide (FeO). (1 MARK)

b In a reaction, 1.45 mol of oxygen was used. How many atoms of oxygen were used in the reaction? (2 MARKS)

Question 18 (5 MARKS)

Argon is sometimes used in light bulbs to stop oxygen from corroding the filament inside the bulb.

a Identify the atomic number for argon. (1 MARK)

b Write the electron configuration for argon. (1 MARK)

c Argon contains three different isotopes: ^{36}Ar, ^{38}Ar and ^{40}Ar. Which isotope of argon would be the most abundant? Explain. (2 MARKS)

d In a 2.44 mol sample of argon, how many atoms of argon would be present? (1 MARK)

Question 19 (6 MARKS)

The gas inhaled prior to surgeries is known as nitrous oxide (N_2O).

a In which period(s) in the periodic table are nitrogen and oxygen found? (1 MARK)

b Which of the elements in nitrous oxide, nitrogen or oxygen, is more electronegative? Explain. (2 MARKS)

c Of the two elements, which is the most reactive? Explain. (2 MARKS)

d If in a reaction, 2.56×10^{24} molecules of nitrous oxide was required, what would this amount be in mol? (1 MARK)

Key science skills

Question 20 (4 MARKS)

Methane is a gaseous fuel that can be found trapped deep under the earth's surface from the decay of organic material. In order to extract methane, excavators need to drill deep into the ground. The gas is then collected in pipes and delivered to the appropriate area.

A group of miners wanted to see if the depth of drilling affected the amount of methane that can be collected. As such, they drilled holes of different depths in what was believed to be a methane rich area.

a Identify the dependent variable. (1 MARK)

b Methane is known to be a light gas. Calculate the number of methane gas molecules that could be collected in 15.8 mol. (1 MARK)

c Describe the key environmental concern related to the collection of methane. (2 MARKS)

Exam-style question hints

14a: A mole of substance contains a fixed number of particles. **14b:** Avogadro's number represents the number of particles/entities in a mole. **15:** To compare values, they need to be in the same units. **16a:** A mole of substance contains a fixed number of molecules. **16b:** A mole of substance contains a fixed number of molecules. **16c:** There can be multiple atoms of the same element in a molecule. **17a:** The formula mass depends on the atoms present. **17b:** A mole of substance contains a fixed number of atoms. **18a:** The atomic number of an element shows its position on the periodic table. **18b:** Energy shells contain subshells. **18c:** The relative atomic mass reflects the relative abundance of each isotope. **18d:** A mole of substance contains a fixed number of atoms. **19a:** The period number depends on the number of energy shells. **19b:** Electronegativity reflects the ability to attract electrons. **19c** Reactivity depends on the behaviour of elements in a reaction. **19d:** A mole of substance contains a fixed number of molecules. **20a:** The dependent variable changes in response to the independent variable. **20b:** A mole of substance contains a fixed number of molecules. **20c:** The process involved in extracting methane affects its surroundings.

2G THE MOLE - PART 2

In this lesson, we will explore how the mole concept is related to the mass of a substance.

2A The atom | 2B Defining elements | 2C The periodic table - part 1 | 2D The periodic table - part 2 | 2E Relative masses | 2F The mole - part 1 | 2G The mole - part 2 | 2H Chemical formulas and equations

Study design dot point

- the mole concept; Avogadro constant; determination of the number of moles of atoms in a sample of known mass; calculation of the molar mass of ionic compounds

Key knowledge units

Molar mass	1.1.15.2
Mass calculations involving the mole	1.1.15.3
Relationship between number and amount of particles	1.1.15.4

Key terms and definitions

Molar mass the mass, in grams, of one mole of any substance

Lesson links

This lesson builds on:

- 2F The mole - part 1

The mole concept can be used when dealing with both the number of particles and the mass of substances.

Molar mass 1.1.15.2

OVERVIEW

The molar mass of each substance is unique.

THEORY DETAILS

In the previous lesson, we found out that the mole quantity was developed based on the number of atoms in 12 g of ^{12}C, which was found to be 6.02×10^{23}. Another important concept that was developed from this idea relates to the mass, in grams, of substances.

As we know, ^{12}C has a relative atomic mass of 12. What we also know is that one mole of ^{12}C atoms has a mass of 12.0 g. The mass of a mole of a substance is known as the **molar mass** and is represented by the symbol M. Based on what we know about ^{12}C, we can see that there is a relationship between the relative atomic mass of an element and the mass of a mole's worth of the same element. If we apply the same thinking to other elements, we will realise that the relationship is the same, as shown in figure 1.

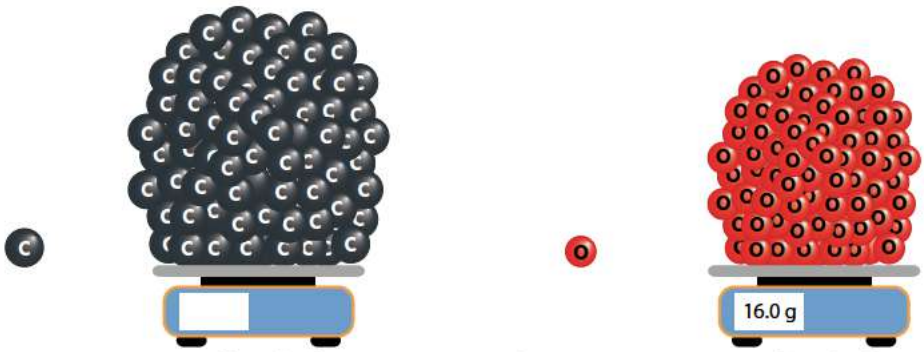


Figure 1 Relationship between the relative atomic mass and molar mass of elements.

As this relationship will always be the same, the molar mass of a substance is related to its atomic mass. Due to the fact that the molar mass represents the mass, in grams, of substance for every mole of a substance, the unit used to represent molar mass is g mol^{-1}.

The molar mass can be calculated for any type of species: elements, compounds and molecules. For example, let's consider carbon monoxide, CO. In order to have one mole of CO, we would need one mole of C atoms and one mole of O atoms. Together, they will be able to make one mole worth of CO molecules as shown in figure 2.

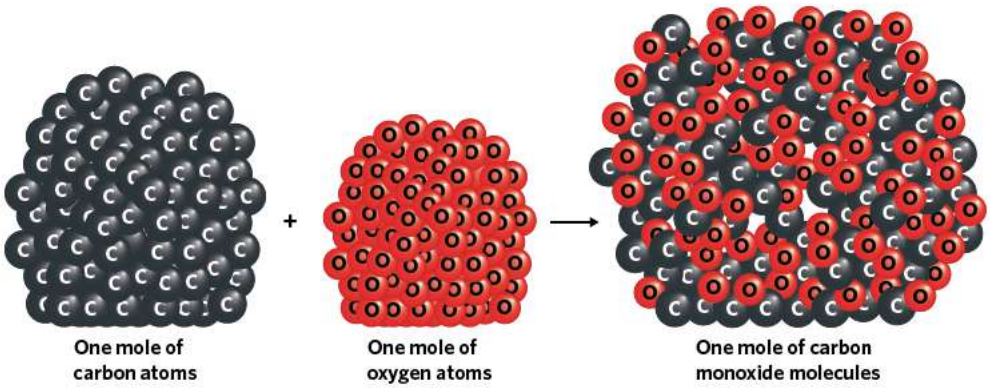


Figure 2 The production of one mole of carbon monoxide, CO.

To calculate the molar mass of one mole of carbon monoxide, we need to add the molar masses of oxygen and carbon together as shown in figure 3.

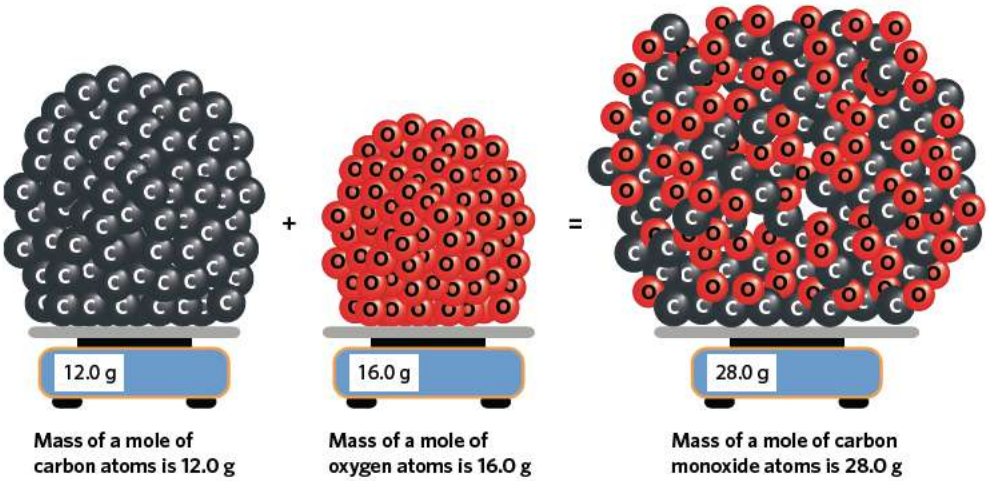


Figure 3 The process of calculating the molar mass of carbon monoxide.

Worked example 1

Nitric oxide, NO, is produced in the body to relax the blood vessels. Calculate the molar mass of nitric oxide.

What information is presented in the question?

The formula for nitric oxide, NO.

What is the question asking us to do?

Calculate the molar mass of NO.

What strategy(ies) do we need in order to answer the question?

1. Use the periodic table to determine the relative atomic mass of each element, N and O.
2. Determine the molar mass of each element.
3. Add the molar mass of both elements together.

Answer

According to the periodic table, the relative atomic mass of N is 14.0 and the relative atomic mass of O is 16.0. Therefore, the molar mass of N and O is 14.0 g mol^{-1} and 16.0 g mol^{-1} respectively.

$M(NO) = 14.0 + 16.0 = 30.0$ g mol^{-1}

Therefore, the molar mass of NO is 30.0 g mol^{-1}.

It is important to note that molecules and compounds frequently contain multiple atoms of the same element. For example, let's consider water. A molecule of water has the molecular formula H_2O, meaning that every molecule of water contains two hydrogen atoms and one oxygen atom. Therefore, in a mole of water molecules, we would need two moles of hydrogen atoms and one mole of oxygen as shown in figure 4.

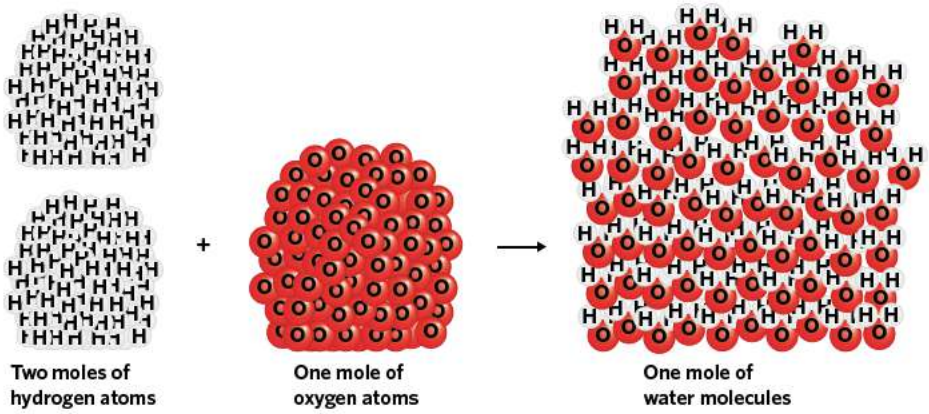


Figure 4 The production of one mole of water, H_2O.

We need to take this into account in our calculations of the molar mass of water as shown in figure 5.

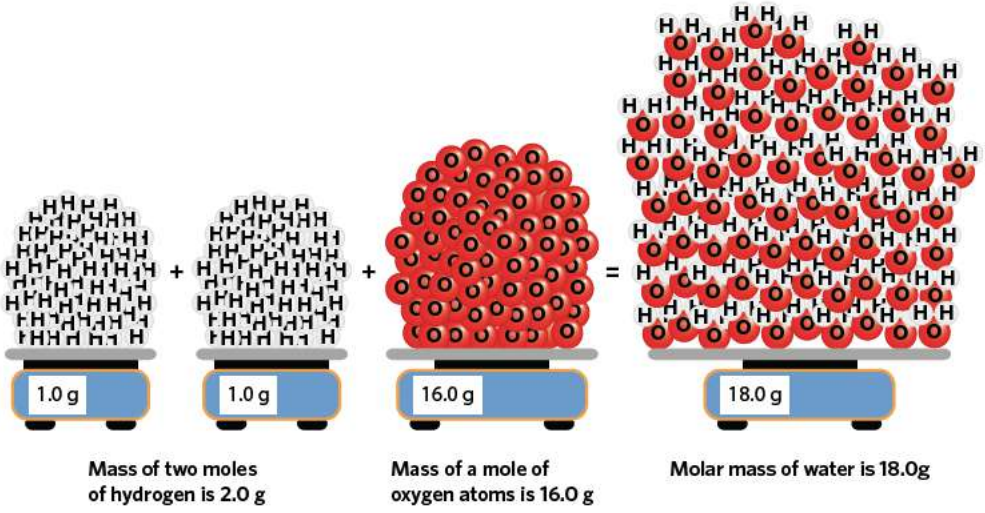


Figure 5 Calculating the molar mass of water.



When calculating the molar mass of compounds involving brackets, the subscript outside of the brackets represents a multiple of what's inside the brackets.

Worked example 2

Copper nitrate, $Cu(NO_3)_2$, is sometimes used as pigments for glass and ceramics. Calculate the molar mass of copper nitrate.

What information is presented in the question?

The formula for copper nitrate, $Cu(NO_3)_2$.

What is the question asking us to do?

Calculate the molar mass of $Cu(NO_3)_2$.

What strategy(ies) do we need in order to answer the question?

1. Use the periodic table to determine the relative atomic mass of each element, Cu, N and O.
2. Determine the molar mass of each element.
3. Add the molar masses of all the elements together.

Answer

According to the periodic table, the relative atomic masses of Cu, N and O are 63.5, 14.0 and 16.0 respectively.
Therefore, the molar masses of Cu, N and O are 63.5 g mol^{-1}, 14.0 g mol^{-1} and 16.0 g mol^{-1} respectively.

$M(Cu(NO_3)_2) = 63.5 + (14 + (16 \times 3)) \times 2 = 187.5$ g mol^{-1}

Hence, the molar mass of copper nitrate is 187.5 g mol^{-1}.

Mass calculations involving the mole 1.1.15.3

OVERVIEW

We can use the molar mass of substances to calculate the amount of substance when given a mass.

THEORY DETAILS

As we have seen above, we can calculate the mass of a mole of a substance. When given a certain mass of a substance, we can determine the amount, in mol, of substance present in the given sample. Let's continue our work with carbon monoxide.

We know that every 28.0 g of carbon monoxide represents one mole of carbon monoxide. Therefore, if we had 56.0 g of carbon monoxide, this would be equivalent to two moles of carbon monoxide. This relationship is represented in figure 6.

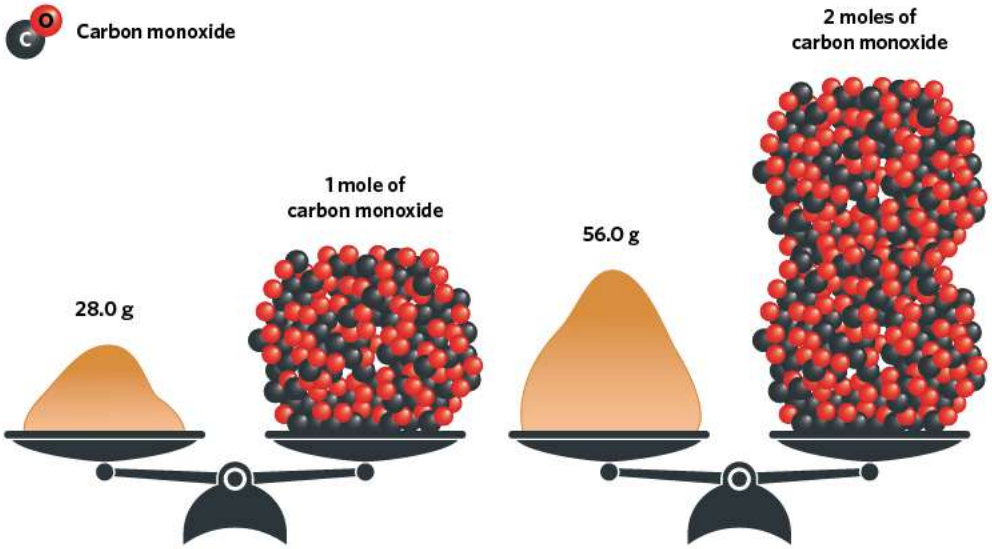


Image: Vector VA/Shutterstock.com

Figure 6 The relationship between mass and amount of moles of carbon monoxide.

Essentially, what we are trying to find out is the number of groups of moles of a substance that can be found in the sample that we have. The way in which we can do this is shown in figure 7. As long as we are able to determine the molar mass of the substance, we will be able to find the amount of substance when given a sample of the substance.

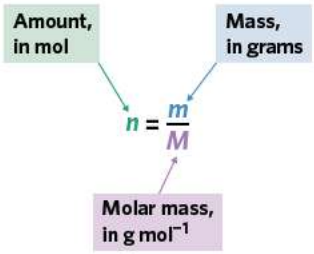


Figure 7 The formula for calculating the amount of substance in mol.

Worked example 3

A student was conducting an experiment that required 12.0 grams of water, H_2O. Calculate the amount of water, in mol, required.

What information is presented in the question?

The mass of water is 12.0 grams.

What is the question asking us to do?

Calculate the amount of water in moles.

What strategy(ies) do we need in order to answer the question?

1. Calculate the molar mass of water.
2. Use the formula $n = \frac{m}{M}$ to determine the amount of water required.

Answer

The molar masses of H and O are 1.0 g mol^{-1} and 16.0 g mol^{-1} respectively. Therefore:

$M(H_2O) = (2 \times 1.0) + 16.0 = 18.0$ g mol^{-1}

To calculate the amount:

$n(H_2O) = \frac{12.0}{18.0} = 0.667$ mol

Hence, 0.667 mol of water is required for the reaction.

Since we know the relationship between the amount of substance and its molar mass, we can use this relationship to determine the mass of a substance. The mass of substance depends on the amount, in mol, of substance present, as shown in figure 8.

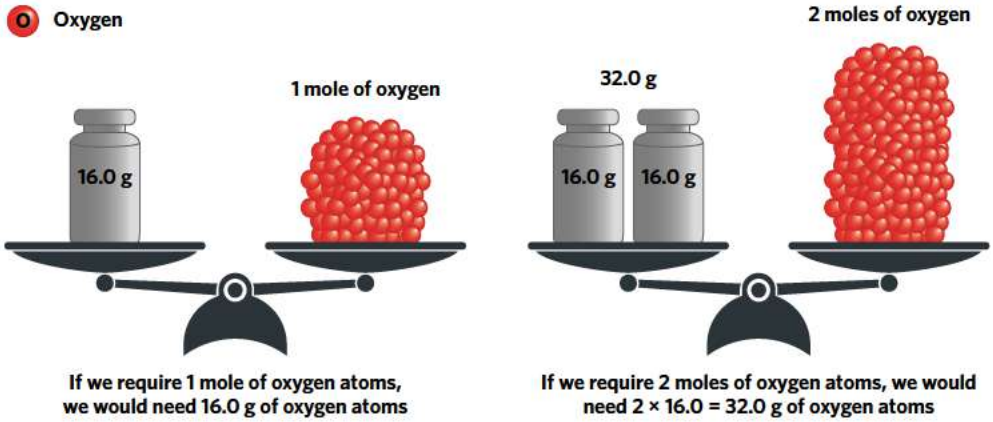


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Figure 8 The relationship between the amount of substance and mass.

If asked to calculate the mass of a substance, in grams, based on the number of moles of substance, all we need to do is multiply the number of moles by the molar mass as shown in figure 9.

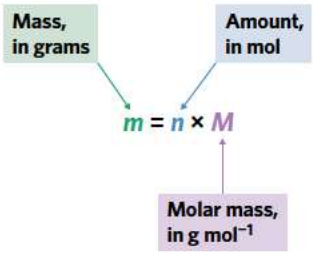


Figure 9 The formula for calculating the mass of a substance.

Worked example 4

Calculate the mass of a 1.66 mol sample of nitrogen dioxide, NO_2.

What information is presented in the question?

The amount of nitrogen dioxide, NO_2, which is 1.66 mol.

What is the question asking us to do?

Calculate the mass of nitrogen dioxide in a 1.66 mol sample.

What strategy(ies) do we need in order to answer the question?

1. Calculate the molar mass of nitrogen dioxide.
2. Use the formula $m = n \times M$ to calculate the mass of nitrogen dioxide.

Answer

The molar masses of N and O are 14.0 g mol^{-1} and 16.0 g mol^{-1} respectively. Therefore:

$M(NO_2) = 14.0 + (2 \times 16.0) = 46.0$ g mol^{-1}

To calculate the mass:

$m(NO_2) = 1.66 \times 46.0 = 76.4$ g

Hence, the mass of nitrogen dioxide, NO_2, in 1.66 mol is 76.4 g.

As we can appreciate, the formulas presented in figures 7 and 8 are different variations of the same formula as they use the same information (variables). This is summarised in figure 10.

We can use the following concept triangle to figure out the type of calculation to perform to find either the amount, in mol, or mass of a substance.

To use this, use your finger to cover up the value you are trying to find. Calculate using the remaining variables as indicated by the signs. For example, to find the amount, cover the section of the triangle that says 'amount', which leaves us to perform the calculation $n = \frac{m}{M}$.

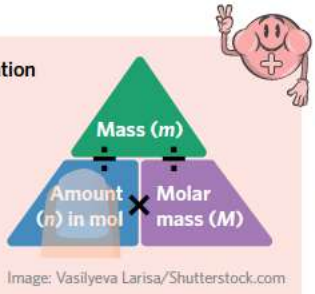


Image: Vasilyeva Larisa/Shutterstock.com

To find the amount, in mol

$$n = \frac{m}{M}$$

To find the mass, in grams

$$m = n \times M$$

Figure 10 Formulas used to calculate the amount in mol and the mass in grams.

Relationship between number and amount of particles 1.1.15.4

OVERVIEW

The mole concept connects the idea of number of particles and mass together.

THEORY DETAILS

In the previous lesson, we learned about the mole in the context of the number of particles present. In this lesson, we explored the mole concept in terms of mass. Based on this, we can appreciate that the mole concept is a bridge between the mass and the number of particles.

This relationship allows us to calculate the number of particles when given mass and vice versa.

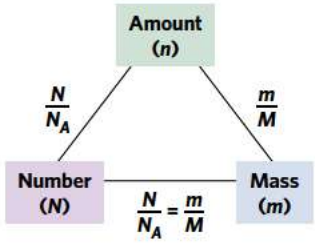


Figure 11 Relationship between the amount, mass and number of particles.

Worked example 5

How many copper atoms are found in 3.98 g of copper?

What information is presented in the question?

The mass of copper which is 3.98 g.

What is the question asking us to do?

Calculate the number of copper atoms in 3.98 g of copper.

What strategy(ies) do we need in order to answer the question?

1. Calculate the amount, in mol, of copper in 3.98 g using the formula $n = \frac{m}{M}$.
2. Calculate the number of copper atoms using $N = n \times N_A$.

Answer

The molar mass of copper is 63.5 g mol^{-1}.

To calculate the mol:

$n(\text{Cu}) = \frac{3.98}{63.5} = 0.06268$ mol

To calculate the number of atoms:

$N(\text{Cu}) = 0.06268 \times 6.02 \times 10^{23} = 3.77 \times 10^{22}$ atoms of copper.

Therefore, there are 3.77×10^{22} atoms of copper in 3.98 g of copper.

Theory summary

- The molar mass is the mass of one mole of substance.
- The amount of substance can be calculated using the mass via $n = \frac{m}{M}$.
- Given that the mole concept is a common variable in both the number of particles and mass, we can calculate the number of particles when given the mass and vice versa.

2G QUESTIONS

Theory review questions

Question 1

The molar mass of an element is equivalent to

A the mass of 6.02×10^{23} atoms of the element.

B the mass of 12 atoms of the element.

Question 2

The molar mass of two different elements

A are the same because there are the same number of particles.

B are different because atoms come in different sizes and masses.

Question 3

The numerical value for the molar mass of an element is the same as

A the number of protons in the element.

B the relative atomic mass of the element.

Question 4

The molar mass of BeO would be calculated by

A adding the molar mass of Be and O.

B subtracting the molar mass of Be and O.

Question 5

The amount, in mol, can be found as long as we have

A the molar mass only.

B both the molar mass and mass.

Question 6

The mass of a substance, in grams, will _____ as the amount of substance decreases.

A increase

B decrease

Question 7

If we are given the mass of a sample of H_2O, we can calculate the amount of H_2O, in mol, by

A multiplying the mass by 18 g mol^{-1}.

B dividing the mass by 18 g mol^{-1}.

Question 8

To find the molar mass of a 1.0 g sample of 0.112 mol of the substance, we would perform the following calculation

A $\frac{1.0}{0.112}$

B $\frac{0.112}{1.0}$

Question 9

Consider the equation $\frac{N}{N_A} = __ = \frac{m}{M}$. The missing symbol is

A M_r.

B *n*.

Question 10

If the number of particles of NO_3 in a substance is known, we can calculate the mass of the NO_3 sample.

A True

B False

Deconstructed questions

Use the following information to answer questions 11–13.

Ammonia, NH_3, is found in many household cleaners. In high concentrations, it can irritate the eyes and skin.

Question 11 (1 MARK)

The symbol N_A represents

A the molar mass of a compound.

B the number of protons in the nucleus.

C 12 particles.

D 6.02×10^{23} particles.

Question 12 (1 MARK)

When given the mass of a substance, the strategy used to find the amount present is

A $n = \frac{M}{m}$

B $n = \frac{M_r}{M}$

C $n = \frac{A_r}{M}$

D $n = \frac{m}{M}$

Question 13 (2 MARKS)

Calculate the mass, in grams, of ammonia (NH_3) present in a sample containing 6.11×10^{26} molecules of ammonia.

Exam-style questions

Within lesson

Question 14 (6 MARKS)

There are millions of different substances that exist due to the large variety of naturally occurring elements on our planet.

a Calculate the molar mass of each of the following compounds.

i Hydroiodic acid, HI. (1 MARK)

ii Magnesium chloride, $MgCl_2$. (1 MARK)

iii Lactic acid, $C_3H_6O_3$. (1 MARK)

b For each pair of compounds, use the greater than or less than symbol to indicate the compound with a greater molar mass.

i $NiCl_2$ __ HNO_3. (1 MARK)

ii H_3PO_4 __ KCl. (1 MARK)

iii KCNS __ HCOONa. (1 MARK)

Question 15 (10 MARKS)

The amount of substance can vary depending on the type of substance.

a For each of the following compounds, calculate the amount, in mol, of substance present.

i 14.9 g of SF_3. (1 MARK)

ii 0.552 g of acetone, C_3H_6O. (1 MARK)

iii 1.12 g of potassium cyanide, KCN. (1 MARK)

b Complete the following calculations to find the quantity of substance present.

i The number of H_2 molecules in 4.60 g of H_2. (2 MARKS)

ii The number of oxygen atoms in 2.10 g of O_2. (3 MARKS)

iii Mass of Be, in grams, given 390 531 atoms of Be. (2 MARKS)

Question 16 (5 MARKS)

Ethanol, C_2H_5OH can be produced from a fermentation process involving glucose, $C_6H_{12}O_6$, and yeast.

$$C_6H_{12}O_{6(aq)} \xrightarrow{\text{yeast}} 2C_2H_5OH_{(aq)} + 2CO_{2(g)}$$

a If 4.11 g of glucose was used to produce ethanol, calculate the amount, in mol, of glucose used. (1 MARK)

b Between ethanol and glucose, which compound has the lowest molar mass? Include calculations in your answer. (2 MARKS)

c What mass of glucose needs to be added so that the student has a total of 3.00 mol of glucose? (2 MARKS)

Multiple lessons

Question 17 (4 MARKS)

In the process of converting fuel into energy, cars can produce carbon monoxide, CO. Carbon monoxide is a poisonous gas that can cause severe health complications.

a If 201.3 mol of carbon monoxide was produced during a car trip, what mass, in kg, of carbon monoxide was produced? (1 MARK)

b Two cars of similar size both claim to have the lowest release of carbon monoxide into the atmosphere. Upon further testing, the results of both cars were recorded in the following table.

	Car A	**Car B**
Quantity of carbon monoxide released per 40 L tank of petrol	114 g	12.04×10^{33} molecules

Which car releases the least amount,in mol, of CO for every tank of fuel? (3 MARKS)

Question 18 (4 MARKS)

The black soot sometimes found on the outside of test tubes that have been placed over a bunsen burner flame is pure carbon.

a Draw a lewis dot diagram for carbon. (1 MARK)

b Write the electronic configuration for carbon. (1 MARK)

c A student wanted to heat 10 g of water in a 5 g test tube. During the experiment, she noticed some black substance forming on the test tube so decided to re-weigh the water-containing test tube. Given that the mass recorded was 17.41 g, calculate the amount, in mol, of carbon formed. (2 MARKS)

Question 19 (5 MARKS)

A cooking recipe displayed quantities of ingredients as shown.

a What mass of sugar was required for this recipe? (1 MARK)

b Calculate the total mass of sugar and bicarbonate in the recipe. (2 MARKS)

c Bicarbonate contains a metallic element. Identify the name of the metallic element and the period in which it would be found on the periodic table. (2 MARKS)

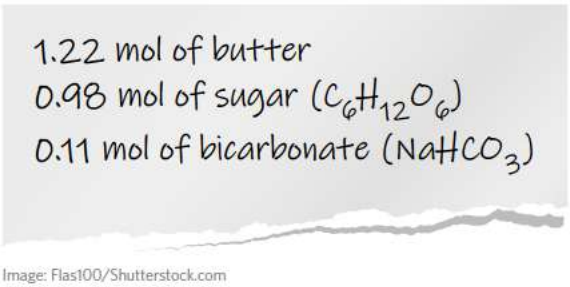


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Key science skills

Question 20 (3 MARKS)

For a science project, Juan was required to produce a scientific poster to illustrate his experiment. For his experiment, he wanted to test the effect of chocolate size on melting time at room temperature.

a Identify the independent variable. (1 MARK)

b Juan ran out of the chocolate he used initially, so decided to substitute it for a different type of chocolate. Comment on how this would affect the reliability of the results. (2 MARKS)

Exam-style question hints

14a: The molar mass includes the mass of all atoms found in the substance. **14b:** Different elements have different molar masses. **15a:** To find the amount of substance when given the mass, we must know the molar mass of the substance. **15b:** *n* is the common variable between the number of particles and the amount of substance. **16a:** To find the amount of substance when given the mass, we must know the molar mass of the substance. **16b:** The molar mass includes the mass of all atoms found in the substance. **16c:** The initial sample may not contain the full amount required. **17a:** The mass of substance depends on the amount of substance. **17b:** Both mass and number of particles can be used to calculate mole. **18a:** Lewis dot diagrams show valence electrons. **18b:** Electronic configurations include subshells. **18c:** The mass of carbon does not include experimental apparatus. **19a:** The mass of a substance depends on the molar mass and amount of substance. **19b:** The total mass includes both substances. **19c:** Metals and non-metals are in different areas on the periodic table. **20a** The independent variable is manipulated by the experimenter. **20b:** Reliability is affected by changes in variables.

2H CHEMICAL FORMULAS AND EQUATIONS

In this lesson, we will learn about the relationship between the mole concept and the development of chemical formulas.

2A The atom	2B Defining elements	2C The periodic table - part 1	2D The periodic table - part 2	2E Relative masses	2F The mole - part 1	2G The mole - part 2	2H Chemical formulas and equations

Study design dot points

- experimental determination of the empirical formula of an ionic compound
- determination of empirical and molecular formulas of organic compounds from percentage composition by mass and molar mass
- the conventions of scientific report writing including chemical terminology and representations, symbols, chemical equations, formulas, units of measurement, significant figures and standard abbreviations

Key knowledge units

Determining empirical formulas	1.2.9.1 & 1.1.16.1
Determining molecular formulas	1.2.9.2
Chemical terminology and representations	2.3.9.1

Key terms and definitions

Percentage composition percentage by mass of an element in a compound

Ratio numerical relationship between two or more elements

Empirical formula simplest, whole number ratio of atoms of different elements in a compound

Balanced equation same number of atoms of each element on either side of the chemical equation

Lesson links

This lesson builds on:

- 2F The mole - part 1
- 2G The mole - part 2

The concept of the mole is important for chemical representations, including molecular formulas, empirical formulas and chemical equations.

Determining empirical formulas 1.2.9.1 & 1.1.16.1

OVERVIEW

Empirical formulas show the simplest whole number ratio of atoms present.

THEORY DETAILS

Compounds are chemical substances that consist of atoms of two or more different elements. Therefore, since compounds may contain different quantities and ratios of each element, chemists have created many different ways to represent molecules, including names, structures and formulas.

Percentage composition by mass

The composition of a compound can be expressed by looking at the percentage of mass that each element in a compound contributes. This is termed **percentage composition** and it shows the proportion of each element present, expressed as a percentage with respect to the mass of the whole compound. For example, methane (CH_4) is a hydrocarbon. The pie chart in figure 1 represents the percentage by mass of carbon and hydrogen atoms in the compound.

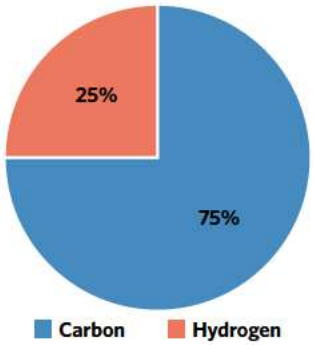


Figure 1 Percentage composition, by mass, of methane (CH_4).

In order to calculate the percentage by mass of an element in a compound, the following formula is used:

$$\%\ by\ mass\ of\ an\ element\ in\ a\ compound = \frac{mass\ of\ element\ in\ 1\ mol\ of\ the\ compound}{molar\ mass\ of\ the\ compound} \times 100$$

Using the example of methane (CH_4) above, in one mol of methane, there is one mol of carbon atoms and therefore, the mass will be:

$m(C) = 1 \times M(C) = 1 \times 12.0 = 12.0$ g

The molar mass of the compound is:

$M(CH_4) = (1 \times 12.0) + (4 \times 1.0) = 16.0$ g mol^{-1}

Using the formula above, calculate the % by mass of carbon in methane:

$$\%\ by\ mass\ of\ carbon\ in\ methane = \frac{12.0}{16.0} \times 100 = 75.0\%$$

In one mol of methane (CH_4), there are four moles of hydrogen atoms and therefore, the mass will be:

$m(H) = 4 \times M(H) = 4 \times 1.0 = 4.0$ g

$$\%\ by\ mass\ of\ hydrogen\ in\ methane = \frac{4.0}{16.0} \times 100 = 25.0\%$$



When calculated correctly, the percentage by mass of each element in a compound should add up to 100%.

Worked example 1

Calculate the percentage by mass of each element in barium sulfate ($BaSO_4$).

What information is presented in the question?

The compound is barium sulfate ($BaSO_4$).

What is the question asking us to do?

Determine the percentage by mass of barium, sulfur and oxygen in the compound.

What strategy(ies) do we need in order to answer the question?

1. Determine the mass of barium, sulfur and oxygen in one mol of the compound.
2. Determine the molar mass of the compound.
3. Use the percentage by mass formula to calculate the percentage composition of each element in the compound.

Answer

In one mol of barium sulfate, there is one mol of barium atoms. Therefore, the mass of barium atoms present in one mol of the compound is:

$m(Ba) = 1 \times M(Ba) = 1 \times 137.3 = 137.3$ g

In one mol of barium sulfate, there is one mol of sulfur atoms. Therefore, the mass of sulfur atoms present in one mol of the compound is:

$m(S) = 1 \times M(S) = 1 \times 32.1 = 32.1$ g

In one mol of barium sulfate, there are four moles of oxygen atoms. Therefore, the mass of oxygen atoms present in one mol of the compound is:

$m(O) = 4 \times M(O) = 4 \times 16.0 = 64.0$ g

The molar mass of the compound is:

$M(BaSO_4) = (1 \times 137.3) + (1 \times 32.1) + (4 \times 16.0) = 233.4$ g mol^{-1}

$$\%\ by\ mass\ of\ barium\ in\ BaSO_4 = \frac{137.3}{233.4} \times 100 = 58.8\%$$

$$\%\ by\ mass\ of\ sulfur\ in\ BaSO_4 = \frac{32.1}{233.4} \times 100 = 13.8\%$$

$$\%\ by\ mass\ of\ oxygen\ in\ BaSO_4 = \frac{64.0}{233.4} \times 100 = 27.4\%$$

Empirical formula

Using the mole concept, atoms of different elements in molecules exist in certain, fixed whole number **ratios**. In chemistry, this ratio represents how many atoms of one element there are compared to another, or even multiple other elements. The simplest, whole number form of this ratio is called the **empirical formula**, where elements are represented by their chemical symbol found on the periodic table, and the ratio with subscript notation. The fact that the ratio is in its simplest form means that any common factors have been cancelled out. For example, if a compound has two carbon atoms and four hydrogen atoms as shown in figure 2, the empirical formula is CH_2 not C_2H_4. This is because the simplest, whole number ratio is 1 : 2, not 2 : 4 because both the number of carbon atoms and the number of hydrogen atoms are divisible by two.

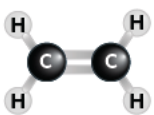
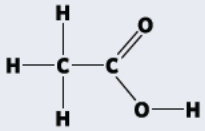
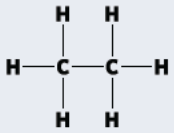
Name	Ethene
Ball-and-stick model	
Empirical formula	CH_2
Ratio	C : H 2 : 4 1 : 2

Figure 2 Empirical formula of ethene and its meaning.

Table 1 illustrates some compounds, their empirical formulas and the simplest, whole number ratio of atoms that they represent.

Table 1 Different compounds and their empirical formulas.

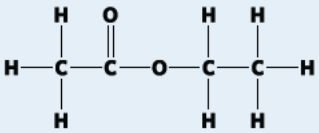
Name and structure	Formula based on the molecular structure	Empirical formula	Simplest, whole number ratio of elements in the compound
Ethanoic acid 	$C_2H_4O_2$	CH_2O	C : H : O 1 : 2 : 1
Ethane 	C_2H_6	CH_3	C : H 1 : 3



Since empirical formulas are the simplest, whole number ratio of elements in a compound, different compounds can have the same empirical formula (e.g. C_5H_{10} and C_6C_{12} both have the empirical formula CH as the simplest, whole number ratio is 1:2).

Worked example 2

What is the empirical formula of the molecule presented?



What information is presented in the question?

The diagram of the molecule.

What is the question asking us to do?

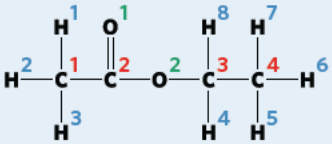
Determine the empirical formula of the compound which is the simplest, whole number ratio of atoms in a compound.

What strategy(ies) do we need in order to answer the question?

1. Identify the elements present.
2. Count the number of atoms of each element present.
3. Express this as a ratio.
4. Determine if a common factor is present which means that the ratio can be further simplified.

Answer

From the picture, there are four carbon atoms, eight hydrogen atoms and two oxygen atoms.



Therefore, the actual ratio of atoms of different elements in the compound is:

C : H : O

4 : 8 : 2

However, 2 is a common factor and therefore, the simplest, whole number ratio is:

C : H : O

$\frac{4}{2} : \frac{8}{2} : \frac{2}{2}$

2 : 4 : 1

Therefore, the empirical formula is C_2H_4O.

Empirical formulas can also be determined from quantitative data. Different types of experiments can determine the mass of each element present in a compound. Using the mole concept, we can convert masses into moles to determine the mole ratio of atoms of particular elements in a compound.



When finding the empirical formula, we only use mole ratios.



Let's take a look at an example. A chemist is investigating an oxide of nitrogen (compound containing only nitrogen and oxygen) which consists of 30.4% nitrogen. She would like to determine the empirical formula of the compound. First of all, since we know that the percentage by mass of elements in a compound must add up to 100%, and the compound contains only nitrogen and oxygen, the percentage by mass of oxygen in the compound is 100% – 30.4% = 69.6%. Secondly, if a compound consists of 30.4% nitrogen by mass and 69.6% oxygen by mass, in 100 g of the compound, there would be 30.4 g of nitrogen and 69.6 g of oxygen. Therefore:

N	:	O
30.4 g	:	69.6 g
$n(\text{N}) = \frac{30.4}{14.0}$	:	$n(\text{O}) = \frac{69.6}{16.0}$
$n(\text{N}) = 2.17$ mol	:	$n(\text{O}) = 4.35$ mol

This means that for every 2.17 mol of nitrogen in the compound, there are 4.35 mol of oxygen. In order to convert this to the empirical formula, we divide both quantities by the lowest amount which in this case is 2.17 mol.

N	:	O
$\frac{2.17}{2.17}$	:	$\frac{4.35}{2.17}$
1.00	:	2.00

Therefore, the empirical formula is NO_2 because the simplest, whole number ratio of nitrogen to oxygen is 1 : 2. This process is shown in figure 3.

Sometimes the mole ratios calculated are not whole numbers. When this happens, we need to multiply the ratio by a factor to find the lowest whole number ratio between the elements.



Percentage is expressed as 'per 100', therefore we can express the % value as a quantity out of 100.

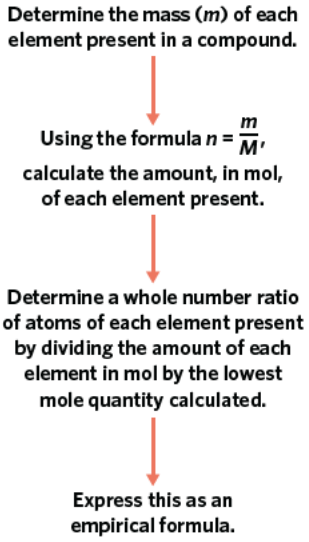


Figure 3 Flowchart of the calculations required to determine the empirical formula.

Worked example 3

Determine the empirical formula of a compound with only carbon, hydrogen and oxygen atoms, consisting of 47.37% carbon and 10.59% hydrogen.

What information is presented in the question?

The compound contains only carbon, hydrogen and oxygen.

It consists of 47.37% carbon and 10.59% hydrogen.

What is the question asking us to do?

To determine the empirical formula of the compound which is the simplest, whole number ratio of atoms in a compound.

What strategy(ies) do we need in order to answer the question?

1. Determine the mass of carbon, hydrogen and oxygen present in 100 g of the compound.
2. Calculate the amount in mol of each element present in 100 g of the compound.
3. Divide each amount by the lowest mole quantity.
4. Determine the simplest, whole number ratio and then express this as an empirical formula.

Answer

Since the compound contains only carbon, hydrogen and oxygen, in 100 g of the compound, there will be 47.37 g of carbon, 10.59 g of hydrogen and

$m(\text{O}) = 100\text{ g} - 47.37\text{ g} - 10.59\text{ g} = 42.04\text{ g}$

Therefore,

C	:	H	:	O
47.37 g	:	10.59 g	:	42.04 g
$\frac{47.37}{12.0} = 3.95$ mol	:	$\frac{10.59}{1.0} = 10.59$ mol	:	$\frac{42.04}{16.0} = 2.63$ mol
3.95 mol	:	10.59 mol	:	2.63 mol
$\frac{3.95}{2.63}$	:	$\frac{10.59}{2.63}$	:	$\frac{2.63}{2.63}$
1.50	:	4.03	:	1.00

Multiply each number by 2 to convert to a whole number ratio.

C : H : O

3 : 8 : 2

Therefore, the empirical formula is $C_3H_8O_2$.

Sometimes, using a table is a helpful way to work through these types of problems. An example of how you can use the table for worked example 3 is shown in table 2.

Table 2 Setting out empirical formula calculations.

	C	H	O
Mass of the element	47.37 g	10.59 g	= 100 − 47.37 − 10.59 = 42.04 g
Amount, in mol, of the element	$\frac{47.37}{12.0} = 3.95$	$\frac{10.59}{1.0} = 10.59$	$\frac{42.04}{16.0} = 2.63$
Divide by the lowest amount, in mol, present	$\frac{3.95}{2.63}$	$\frac{10.59}{2.63}$	$\frac{2.63}{2.63}$
Ratio of elements	1.50	4.03	1.00
Lowest whole number ratio	2 × 1.50 = 3	2 × 4.03 = 8	2 × 1.00 = 2

Worked example 4

Polyethylene terephthalate (PET) is a compound containing only carbon, hydrogen and oxygen found in plastic bottles. 3.80 g of the compound was burnt in oxygen which produced 1.32 g of H_2O and 8.04 g of CO_2. What is the empirical formula of the compound?

What information is presented in the question?

The compound consists of only carbon, hydrogen and oxygen.

Mass of compound burnt: 3.80 g

Mass of H_2O produced: 1.32 g

Mass of CO_2 produced: 8.04 g

What is the question asking us to do?

Determine the empirical formula of the compound which is the simplest, whole number ratio of atoms in a compound.

What strategy(ies) do we need in order to answer the question?

1. Determine the amount, in mol, of hydrogen present in the compound.
2. Determine the amount, in mol, of carbon present in the compound.
3. Determine the mass of carbon and hydrogen in the compound.
4. Determine the mass of oxygen in the compound.
5. Determine the amount, in mol, of oxygen in the compound.
6. Divide each amount in mol by the lowest amount to determine the simplest ratio of elements in the compound.
7. Express this ratio as a whole number and in empirical formula notation.

Answer

$n(H_2O) = \frac{1.32}{(2 \times 1.0) + 16.0} = 0.073$ mol

$n(H) = 2 \times n(H_2O) = 0.15$ mol

$m(H)$ *in compound* $= 0.15 \times M(H) = 0.15 \times 1.0 = 0.15$ g

$n(CO_2) = \frac{8.04}{12.0 + (2 \times 16.0)} = 0.18$ mol

$n(C) = 1 \times n(CO_2) = 0.18$ mol

$m(C)$ *in compound* $= 0.18 \times M(C) = 0.18 \times 12.0 = 2.19$ g

$m(O)$ *in compound = total mass of compound* $- m(C) - m(O)$

$= 3.80 - 2.19 - 0.15 = 1.46$ g

$n(O) = \frac{1.46}{16.0} = 0.091$ mol

C	:	H	:	O
0.18 mol	:	0.15 mol	:	0.091 mol
$\frac{0.18}{0.091}$	:	$\frac{0.15}{0.091}$	:	$\frac{0.091}{0.091}$
1.9	:	1.6	:	1.0

Multiply each number by 2 to convert to a whole number ratio.

C : H : O

4 : 3 : 2

Therefore, the empirical formula is $C_4H_3O_2$.

Determining molecular formulas 1.2.9.2

OVERVIEW

Molecular formulas depict the actual number of atoms present in a compound.

THEORY DETAILS

While the empirical formula is the simplest, whole number ratio of atoms of each element in a compound, the molecular formula depicts the actual number of each type of element present in the compound as shown in figure 4.

As we can see in figure 4, the molecular formula C_2H_6, indicates that two atoms of the element represented by the symbol C and six atoms of the element represented by symbol H are present. This can then be translated, using the periodic table and the chemical symbols represented on it, as two atoms of carbon and six atoms of hydrogen. However, the empirical formula is CH_3 because the simplest, whole number ratio of carbon atoms to hydrogen atoms is 1 : 3.

In this example, the molecular formula and empirical formula were different. However, this is not always the case. For example, water's molecular formula is H_2O because there are two atoms of hydrogen and one atom of oxygen in each molecule. Furthermore, the simplest, whole number ratio of hydrogen atoms to oxygen atoms is 2 : 1 and therefore, water's empirical formula is also H_2O.

The molecular formula of a compound can be determined from its empirical formula if the molar mass of the compound is known. This is because the molar mass of a compound's molecular formula will always be some whole number multiple of the molar mass of the empirical formula.

The number that multiplies the empirical formula to get to the molecular formula is calculated by the following equation:

$$empirical\,formula\,units = \frac{molar\ mass\ of\ the\ compound\ (molar\ mass\ of\ the\ molecular\ formula)}{molar\ mass\ of\ the\ empirical\ formula}$$

The process to determine the molecular formula of a compound from its empirical formula is shown in figure 5.

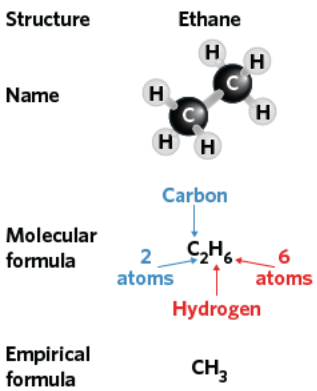


Figure 4 A breakdown of the molecular formula of ethane.



Sometimes a molecule's empirical formula is the same as its molecular formula.

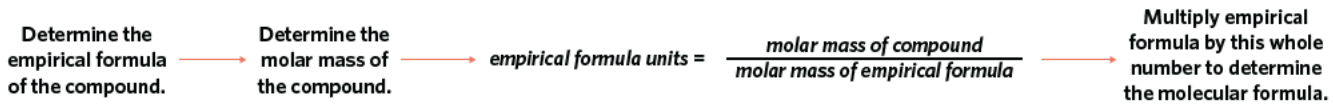


Figure 5 Flowchart of the process required to determine the molecular formula of a compound from an empirical formula.

Worked example 5

The empirical formula of a compound is CH_2O. Through a technique known as mass spectrometry, the molar mass of the compound is shown to be 60 g mol^{-1}. What is the compound's molecular formula?

What information is presented in the question?

Empirical formula: CH_2O

Molar mass: 60 g mol^{-1}

What is the question asking us to do?

Determine the molecular formula.

What strategy(ies) do we need in order to answer the question?

1. Determine the molar mass of the empirical formula.
2. Determine the number of empirical formula units in the molar mass of the compound.
3. Multiply the empirical formula by the number of empirical formula units to determine the molecular formula.

Answer

Molar mass of empirical formula =
$1 \times 12.0 + 2 \times 1.0 + 1 \times 16.0 = 30.0$ g mol^{-1}

Empirical formula units $= \frac{60.0}{30.0} = 2$

Molecular formula is $2 \times (CH_2O)$

$= C_{1\times2}H_{2\times2}O_{1\times2} = C_2H_4O_2$

Chemical terminology and representations 2.3.9.1

OVERVIEW

Just like mathematics, chemistry has its own language with different symbols, representations and terminology.

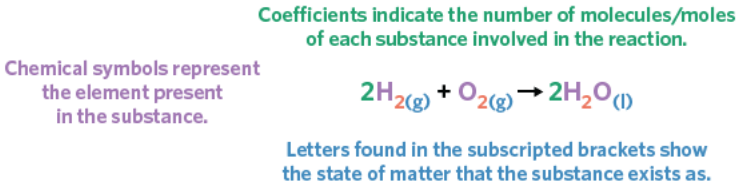
THEORY DETAILS

An important part of chemistry is investigating chemical reactions. Chemical reactions involve the rearrangement of atoms found in the reactants to produce different products. This process is illustrated by a chemical equation as shown.

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

Reactants → Products

Chemical equations show a lot of information about a chemical reaction as shown in figure 6.





If there's isn't a coefficient present, imagine that there is a 1 in front of the formula.

Figure 6 Key elements in a chemical equation.

As we know, substances can exist in different states, each of which have their own convention when represented in a chemical reaction as shown in table 3.

Table 3 States of matter conventions.

Convention	State
(g)	Gas
(s)	Solid
(l)	Liquid
(aq)	Aqueous (chemical species dissolved in water, also known as as a solution)

Another important component of a chemical equation is the use of coefficients in front of the formula of each substance. These numbers represent the ratio between each of the substances present, and links to the mole concept. This is an important concept that we will continue to revisit throughout our studies.

In the chemical reaction shown, we can see that hydrogen gas reacts with oxygen gas to produce liquid water. The product (H_2O) formed is made from atoms that were present in the reactant molecules (H_2 and O_2). If we look even closer, we will notice that not only are the elements present in the reactants and products the same, but the number of atoms of each element is also the same. When this is the case, we can say that the chemical equation is **balanced**. This highlights a concept known as the conservation of matter, whereby the total number and type of atoms in a system remains the same before and after a chemical change, where atoms are neither created nor destroyed.

To see if a chemical equation is balanced, we do a quick check as shown in figure 7 and table 4.

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

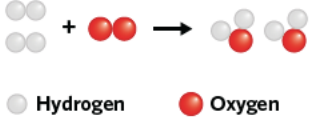


Figure 7 Process involved to check if a chemical equation is balanced.

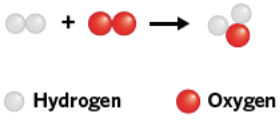
Table 4 Process of ensuring that reactants and products are balanced in a chemical equation.

	Reactants	Products
Hydrogen	There are two hydrogen atoms per molecule, however we have a total of two hydrogen molecules as indicated by the coefficient 2. Therefore, we have a total of: 2 × 2 = 4 hydrogen atoms	Hydrogen is found as part of the water molecule. Each water molecule contains two hydrogen atoms. The coefficient in front of the H_2O in the equation shows that there are two water molecules in total. Therefore, we have a total of: 2 × 2 = 4 hydrogen atoms
Oxygen	There are two oxygen atoms per molecule. We only have one molecule of oxygen and therefore we have a total of: 1 × 2 = 2 oxygen atoms	Oxygen is found as part of the water molecule. Each water molecule contains one oxygen atom. The coefficient in front of the H_2O in the equation shows that there are two water molecules in total. Therefore, we have a total of: 2 × 1 = 2 oxygen atoms



The concept of a balanced chemical equation is extremely important as it is an underlying skill that is required for all of the chemical theory that we will be learning about. Therefore, it's important that we understand how to balance an equation.

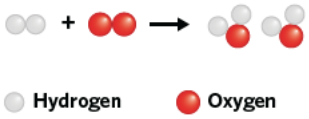
Going back to the original chemical reaction presented earlier between hydrogen gas and oxygen gas, the initial equation would be as shown in figure 8.



	Present in the reactant side	Present in the product side
Hydrogen	Two	Two
Oxygen	Two	One

Figure 8 Initial equation for the reaction between hydrogen and oxygen.

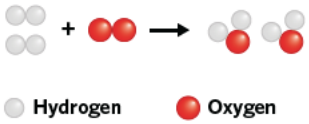
The number of hydrogen atoms found on the reactant and product sides is the same, however there is an imbalance of oxygen atoms, with a greater number of oxygen atoms on the reactant side than the product side. To balance the oxygen atoms, we need to add another oxygen atom on the product side. However, we cannot add single atoms. Since oxygen is present as part of a water molecule, in order to add another oxygen atom, we must add a whole water molecule to the reaction. Therefore, the equation would change as shown in figure 9.



	Present in the reactant side	Present in the product side
Hydrogen	Two	Four
Oxygen	Two	Two

Figure 9 Balancing of oxygen atoms in a reaction between hydrogen and oxygen.

By adding the additional water molecule, we have also added another two hydrogen atoms to the product side, which has caused an imbalance in the number of hydrogen atoms. To balance the number of hydrogen atoms, we need more hydrogen atoms on the reactant side. To do so, we can add a hydrogen molecule so that the equation changes as shown in figure 10.



	Present in the reactant side	Present in the product side
Hydrogen	Four	Four
Oxygen	Two	Two

Figure 10 Balancing of hydrogen atoms in a reaction between hydrogen and oxygen.

The equation is now balanced. As we can see, in the balanced equation, there are two hydrogen gas molecules, one oxygen gas molecule and two water molecules. Therefore the balanced equation is written as:

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

According to the balanced equation, for every two molecules of hydrogen that reacts with one molecule of oxygen, two molecules of water can be formed.



Do not add single atoms and change the subscript of elements in formulas when balancing equations. This changes the substance entirely!

Worked example 6

Balance the following chemical equation which represents the process of photosynthesis. Interpret what the numbers in front of the molecular formulas mean.

$CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + O_{2(g)}$

What information is presented in the question?

Unbalanced chemical equation for photosynthesis:

$CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + O_{2(g)}$

What is the question asking us to do?

Balance the chemical equation.

What strategy(ies) do we need in order to answer the question?

1. Determine the number of carbon, hydrogen and oxygen atoms on either side of the equation.
2. Change the number of atoms on either side of the equation by adding molecules where appropriate, which changes the coefficients in the equation.

Answer

Initially, there is one carbon atom, two hydrogen atoms and three oxygen atoms on the reactant side of the equation and six carbon atoms, 12 hydrogen atoms and eight oxygen atoms on the product side of the equation.

In order to balance the carbon atoms, we need a total of six CO_2 molecules:

$6CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + O_{2(g)}$

There are now:

Reactant	:	Product
6 carbon atoms	:	6 carbon atoms
13 oxygen atoms	:	8 oxygen atoms
2 hydrogen atoms	:	12 hydrogen atoms

In order to balance the hydrogen atoms, we need six H_2O molecules:

$6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + O_{2(g)}$

There are now:

Reactant	:	Product
6 carbon atoms	:	6 carbon atoms
18 oxygen atoms	:	8 oxygen atoms
12 hydrogen atoms	:	12 hydrogen atoms

In order to balance the oxygen atoms, we need six O_2 molecules:

$6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + 6O_{2(g)}$

There are now:

Reactant	:	Product
6 carbon atoms	:	6 carbon atoms
18 oxygen atoms	:	18 oxygen atoms
12 hydrogen atoms	:	12 hydrogen atoms

The equation is now balanced.

The numbers in front of the molecular formulas indicate the ratio of different molecules reacting. In this case, six molecules of CO_2 react with six molecules of H_2O to form one molecule of $C_6H_{12}O_6$ and six molecules of O_2.

Given that balanced chemical equations represent the number of molecules required in a reaction and the number of molecules produced in a reaction, they also represent mole ratios. If we recall lesson 2F, the mole is a unit which represents a fixed number of molecules/atoms (Avogadro's number). Therefore, using the example above, the mole ratio is:

$6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + 6O_{2(g)}$

CO_2	:	H_2O	:	$C_6H_{12}O_6$	:	O_2
6	:	6	:	1	:	6

This means that six mol of CO_2 reacts with six mol of H_2O to produce one mol of $C_6H_{12}O_6$ and six mol of O_2. Based on this ratio, we can calculate the amount required for a specific reaction to occur, or the amount of product formed. The general formula to perform these calculations is:

$$n(unknown) = \frac{WANT}{KNOW} \times n(known)$$

where WANT is the coefficient or number in front of the chemical species whose amount we want to determine and KNOW is the coefficient or number in front of the chemical species whose amount we know. Using the example above, we can now determine the amount of $C_6H_{12}O_{6(aq)}$ produced. For example, if one mol of $CO_{2(g)}$ reacts with sufficient $H_2O_{(l)}$, 0.17 mol of $C_6H_{12}O_6$ will be produced as shown in figure 11.

$6CO_{2(g)} + 6H_2O_{(l)} \longrightarrow 1C_6H_{12}O_{6(aq)} + 6O_{2(g)}$

KNOW WANT

$n(C_6H_{12}O_6) = \frac{1}{6} \times 1 = 0.17$ mol

Figure 11 Steps to determine the amount of an unknown substance from a substance where the amount is known.



When balancing complicated equations, it may be easier to start by balancing key elements; those that are only found in one product.

Theory summary

- Percentage composition by mass reveals the percentage that the atom(s) of each element contribute to the total mass of the compound.
- Empirical formulas reveal the simplest, whole-number ratio of atom(s) of each element in a compound.
- Molecular formulas reveal the actual number of atom(s) of each element in a compound.
- Chemical equations when balanced have the same number of atom(s) of each element on either side of the reaction.
- A balanced equation shows the ratios of the molecules / amount of substance (in mol) involved in a chemical reaction.

2H QUESTIONS

Theory review questions

Question 1

A compound consists of only carbon and hydrogen. If the percentage by mass of carbon is 75%, what is the percentage by mass of hydrogen?

A 25%

B 75%

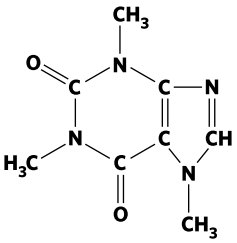
Question 2

What is the percentage by mass of carbon in C_2H_5OH?

A 26%

B 52%

The following diagram applies to questions 3–5.



Question 3

What is the molecular formula of the pictured compound?

A $C_8H_{10}N_4O_2$

B $C_4H_5N_2O$

Question 4

What is the empirical formula of the pictured compound?

A $C_8H_{10}N_4O_2$

B $C_4H_5N_2O$

Question 5

Empirical and molecular formulas are always different.

A True

B False

Question 6

The molecular formula of a compound is known. What piece of information is required to determine the empirical formula?

A The molar mass of the compound.

B No other piece of information is required.

Question 7

The empirical formula of a compound is known. What piece of information is required to determine the molecular formula of the compound?

A The molar mass of the compound.

B No other piece of information is required.

Question 8

Is the following chemical equation balanced?

$Na_{(s)} + O_{2(g)} \rightarrow Na_2O_{(s)}$

A Yes

B No

Question 9

In order to make this equation balanced, what are the values of *x* and *y* respectively?

$xAl_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + yFe_{(s)}$

A 2, 1

B 2, 2

Question 10

What does the following chemical equation mean?

$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

A One molecule of $Mg_{(s)}$ reacts with two molecules of $HCl_{(aq)}$.

B One molecule of $Mg_{(s)}$ reacts with one molecule of $H_2Cl_{2(aq)}$.

Deconstructed questions

Use the following information to answer questions 11–13.

Iron oxides are chemical compounds which are present in large quantities in Australia. A chemist working for a mining company was investigating the properties of these compounds. She burnt 0.424 g of iron in the presence of excess oxygen and produced 0.606 g of a reddish-brown oxide.

Question 11 (1 MARK)

What is the amount of iron, in mol, present in the compound?

A 0.424 mol

B 55.8 mol

C 0.00760 mol

D 0.0109 mol

Question 12 (1 MARK)

What is the amount of oxygen, in mol, present in the compound?

A 0.0379 mol

B 0.0265 mol

C 0.182 mol

D 0.0114 mol

Question 13 (3 MARKS)

Based on the above answers, determine the empirical formula of the compound. The chemist also discovered that the molar mass of the compound is 159.6 g mol^{-1}. Is the empirical formula the same as the molecular formula of the compound? Justify your answer.

Exam-style questions

Within lesson

Question 14 (13 MARKS)

For the following compounds, calculate the percentage by mass of

a H in HCl (2 MARKS)

b K in $KMnO_4$ (2 MARKS)

c Mn in $KMnO_4$ (2 MARKS)

d N in $(NH_4)_3PO_4$ (2 MARKS)

e H in $(NH_4)_3PO_4$ (2 MARKS)

f Fe in Fe_2O_3. What is the mass of iron that can be recovered from a 36.2 g sample of Fe_2O_3? (3 MARKS)

Question 15 (10 MARKS)

Based on the following pieces of information, calculate the empirical formula of the compounds given.

a H: 2.8% Cl: 97.2% (2 MARKS)

b Pb: 38.43 g C: 17.83 g H: 3.74 g (2 MARKS)

c Al: 15.8% S: 28.1% O: 56.1% (2 MARKS)

d A 6.4 g compound containing hydrogen and carbon that contains 4.8 g of carbon. (2 MARKS)

e A 1.996 g compound composed of only nitrogen and oxygen contains 0.608 g of nitrogen. (2 MARKS)

Question 16 (10 MARKS)

Based on the following pieces of information, determine the molecular formula of the compounds given.

a Empirical formula: CH_2 molar mass: 70.0 g mol^{-1}. (2 MARKS)

b Empirical formula: CH_2O molar mass: 180.0 g mol^{-1}. (2 MARKS)

c Empirical formula: $C_6H_{10}S_2O$ molar mass: 162.2 g mol^{-1}. (2 MARKS)

d Empirical formula: CH_4N molar mass: 60.0 g mol^{-1}. (2 MARKS)

e Empirical formula: C_3H_8N molar mass: 116.0 g mol^{-1}. (2 MARKS)

Question 17 (12 MARKS)

For each of the following unbalanced chemical equations:

i Balance the chemical equation.

ii Explain in words the mole ratio between products and reactants.

iii For equations *b* and *c*, calculate the amount, in mol, of reactants required to produce one mol of product.

a $CH_{4(g)} + Cl_{2(g)} \rightarrow CCl_{4(g)} + HCl_{(g)}$ (2 MARKS)

b $Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$ (3 MARKS)

c $Al_{(s)} + Cl_{2(g)} \rightarrow AlCl_{3(g)}$ (3 MARKS)

d $SeCl_{6(g)} + O_{2(g)} \rightarrow SeO_{2(s)} + Cl_{2(g)}$ (2 MARKS)

e $H_2O_{(l)} + CO_{2(g)} \rightarrow C_7H_{8(l)} + O_{2(g)}$ (2 MARKS)

Multiple lessons

Question 18 (8 MARKS)

Lithium ($Li_{(s)}$) is a highly reactive metal and in the presence of water ($H_2O_{(l)}$), it will spontaneously react to produce lithium hydroxide ($LiOH_{(aq)}$) and flammable hydrogen gas ($H_{2(g)}$).

a Write the unbalanced chemical equation for the reaction between lithium and water. (1 MARK)

b Balance this chemical equation. Explain your reasoning. (2 MARKS)

c What is the electron configuration of lithium according to Schrödinger's model of the atom? (1 MARKS)

d Compare the electronegativity of lithium and fluorine. (2 MARKS)

e Compare the reactivity of lithium and potassium. (2 MARKS)

Question 19 (10 MARKS)

In today's society, lots of energy is produced by the complete combustion of hydrocarbons. The general combustion reaction of a hydrocarbon involves the hydrocarbon C_aH_b, where a and b are constants reacting with oxygen, $O_{2(g)}$, to produce water ($H_2O_{(g)}$) and carbon dioxide ($CO_{2(g)}$).

a A 0.481 g sample of hydrocarbon is burnt. It produces 1.24 g of CO_2 and 0.255 g of H_2O. What is the hydrocarbon's empirical formula? (3 MARKS)

b If the molar mass of the compound is 78.0 g mol^{-1}, what is the hydrocarbon's molecular formula? (2 MARKS)

c Write the unbalanced chemical equation for the complete combustion of this hydrocarbon. (1 MARK)

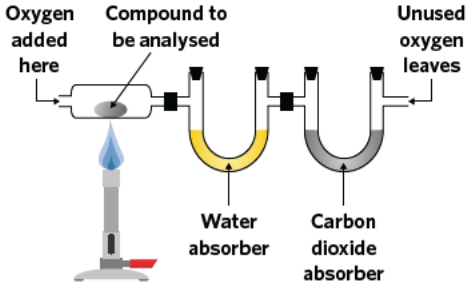
d Balance this chemical equation. Justify your answer. (2 MARKS)

e Which element, carbon or oxygen, has the greater electronegativity? Justify your answer. (2 MARKS)

Key science skills

Question 20 (6 MARKS)

The apparatus shown is used in combustion analysis to determine the empirical formula of compounds containing carbon, hydrogen and oxygen.



a The experimenter unseals the water absorber, exposing it to the atmosphere, but leaves the carbon dioxide absorber sealed. Comment on the accuracy of the results. (2 MARKS)

b The scale that is used is not calibrated. Identify the type of error. (2 MARKS)

c The experimenter says that due to the uncalibrated scale, the results won't be precise but will be accurate. Is this correct? (2 MARKS)

Exam-style question hints

14: Percentage by mass is the percentage that an element contributes to the total mass of the compound. **15:** The empirical formula is the simplest, whole-number ratio of atoms of each element present. **16:** The empirical formula and molecular formula are multiples of each other. **17:** Balanced chemical equations have the same number of atoms of each element on either side of the reaction arrow. **18a:** The reactants are on the left-hand side of the chemical equation and the products are on the right-hand side. **18b:** Balanced chemical equations have the same number of atoms of each element on each side of the chemical equation. **18c:** Schrödinger's model of the atom involves shells and subshells. **18d:** Electronegativity is a measure of how strongly an atom attracts electrons to itself. **18e:** Reactivity for Group 1 elements is how easily they lose their valence electron. **19a:** The empirical formula is the simplest, whole-number ratio of atoms of each element present. **19b:** The empirical formula and molecular formula are multiples of each other. **19c:** The reactants are on the left-hand side of the chemical equation and the products are on the right-hand side. **19d:** Balanced chemical equations have the same number of atoms of each element on each side of the chemical equation. **19e:** Electronegativity is a measure of how strongly an atom attracts electrons to itself. **20a:** Accuracy is a measure of how close the measured value is to the true value. **20b:** Different errors have different effects on the results. **20c:** Precision is a measure of how close measured values are to each other, not the true value.

EXPERIMENT

DETERMINATION OF THE AVERAGE ATOMIC MASS OF A NEWLY DISCOVERED ELEMENT

Scientists have recently ventured beyond the known universe to a planet known by its sweet local inhabitants as Sugarurn. They have discovered an element called M&Mine and so would like to determine its average atomic mass. It is known that the element consists of peanut, almond and sultana isotopes.

Materials

(this will be enough for two–three groups to use)

- 1 × Packet of chocolate coated sultanas
- 1 × Packet of chocolate coated peanuts
- 1 × Large tub
- 2 × Packets of chocolate coated almonds
- 1 × Electronic balance
- 4 × Plastic cups

Note: the chocolates can be substituted for any other type of food that exists in different varieties.

Method

1. Open all of the packets of chocolates and pour them into the tub. Mix well.
2. Using one plastic cup, scoop out a full cup of chocolate.
3. Label each of the remaining cups A, B and C.
4. Measure the mass of each cup (A, B and C).
5. Separate each chocolate type (isotopes) into the three different plastic cups labelled A, B and C.
6. Use the balance to record the mass of cup A (including the chocolates).
7. Count the number of chocolates in cup A.
8. Divide the mass of cup A by the number of chocolates in cup A to determine the average mass of each chocolate found in cup A.
9. Repeat steps 6–8 for cups B and C.
10. Determine the percentage of each isotope in the element by dividing the number of chocolates in each cup by the total number of chocolates and multiplying by 100.
11. Calculate the average atomic mass of M&Mine.

Results

	Cup A	Cup B	Cup C
Mass of the cup (g)			
Mass of the cup + chocolates (g)			
Number of chocolates in the cup			
Average mass of each chocolate (g)			

QUESTIONS

Question 1 (1 MARK)

What type of data is being collected?

Question 2 (2 MARKS)

Why were the chocolates mixed?

Question 3 (2 MARKS)

Why was the whole cup of chocolates weighed and then divided by the number of individual chocolates found in the cup as opposed to weighing a single chocolate to determine the average mass of each chocolate?

Question 4 (4 MARKS)

The following is an excerpt from a student's calculations:

Average mass of sultana chocolate isotope = 1.02 g

Average mass of almond isotope = 1.13 g

Average mass of peanut isotope = 1.05 g

Abundance of sultana chocolate isotope = 49%

Abundance of almond isotope = 30%

Abundance of peanut isotope = 21%

$$\text{Average atomic mass} = \frac{(1.02 + 1.13 + 1.05)}{3} = 1.07 \text{ g}$$

a Identify the error in the student's calculations. What type of error is this? (2 MARKS)

b Calculate the true average atomic mass of M&Mine. (2 MARKS)

Question 5 (2 MARKS)

A student decides to eat some of the M&Mine. Identify two possible safety concerns.

Question 6 (2 MARKS)

Calculate the average atomic mass of M&Mine using the data collected in your experiment.

ANSWERS

1 Quantitative data

2 [M&Mine is a mixture of different isotopes.[1]][Therefore, in order to get a sample in a plastic cup that is representative of the element's true percentage composition, the chocolates need to be thoroughly mixed so their abundance is equivalent throughout the mixture.[2]]

✓ ✗ I have identified M&Mine as a mixture of different isotopes.[1]

✓ ✗ I have explained the need to get an accurate sample.[2]

3 [Weighing a larger mass on an electronic balance reduces the likelihood of random errors arising from an incorrect reading.[1]][This improves the accuracy of the experiment because not only are electronic balances better suited for larger masses, the average mass of a chocolate can be calculated as opposed to the single mass of one chocolate.[2]]

✓ ✗ I have explained the reason behind weighing a larger mass.[1]

✓ ✗ I have justified my answer with reference to the accuracy of the experiment.[2]

4 **a** [The error is the student has not taken into account the relative abundances of each isotope when calculating the average atomic mass of the element.[1]][This is a personal error because it was a miscalculation on the student's behalf.[2]]

✓ ✗ I have identified the error in the student's calculations.[1]

✓ ✗ I have identified the type of error.[2]

b $$\text{Average atomic mass} = \frac{(1.02 \times 49) + (1.13 \times 30) + (1.05 \times 21)}{100} = 1.06 \text{ g}$$

5 [Since the M&Mine is being touched by many different hands, there is the possibility of contamination with bacteria.[1]]
[Furthermore, the laboratory has many dangerous chemicals and so the M&Mine could have become contaminated with chemicals.[2]]

✓ ✗ I have identified the risk of infection spreading.[1]

✓ ✗ I have identified the risk of chemical contamination.[2]

6 *Please check your answer with your teacher.

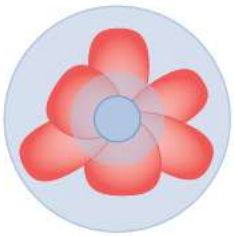
CHAPTER 2 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

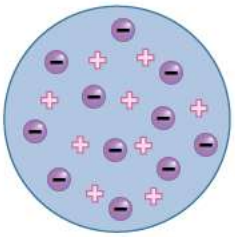
Question 1 (1 MARK)

Which of the following diagrams is representative of Bohr's model of the atom?

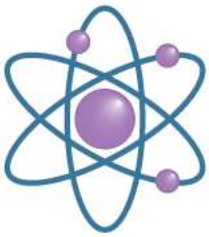
A



B



C



D



Question 2 (1 MARK)

What is the electronic configuration according to Schrödinger's model of the atom in Period 3, Group II of the periodic table?

A $1s^2 2s^2 2p^6 3s^2$

B $1s^2 2s^2 2p^6$

C 2,8,2

D 2,3

Question 3 (1 MARK)

Which of the following statements regarding atomic number is incorrect?

A Atomic numbers are unique to each element.

B The atomic number represents the number of electrons in a neutrally charged atom.

C The atomic number is the number of neutrally charged particles in the nucleus.

D The periodic table is organised with respect to atomic number.

Question 4 (1 MARK)

Which of the following equations is always true where A = atomic number, P = number of protons, N = number of neutrons, and M = mass number?

A $M = 2 \times P$

B $M = A + P$

C $M = 2 \times N$

D $M = A + N$

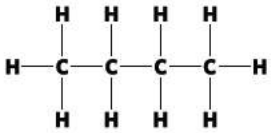
Question 5 (1 MARK)

Which of the following options correctly matches the properties of elements to their trend in the periodic table?

	Property	Across a period (left to right)	Down a group
A	Electronegativity	Increases	Increases
B	Core charge	Increases	Decreases
C	First ionisation energy	Decreases	Increases
D	Metallic character	Decreases	Increases

Question 6 (1 MARK)

Butane is the molecule pictured and it is often used as a fuel for cooking while camping.

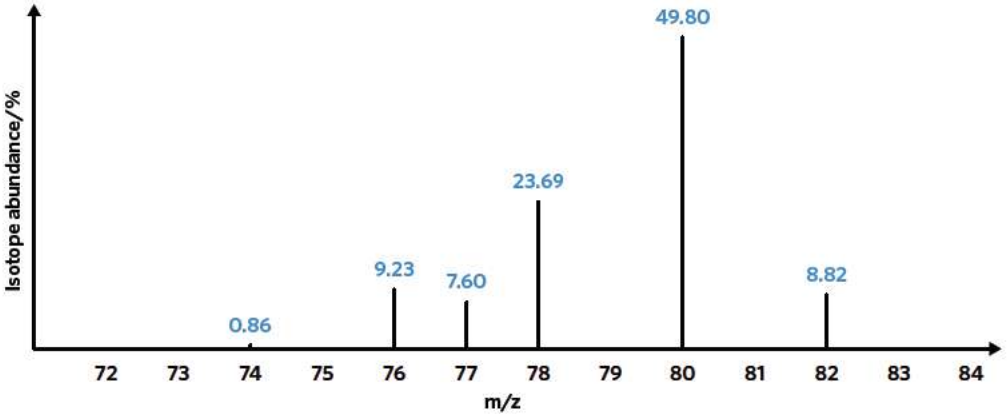


What is the relative molecular mass, empirical formula and molecular formula of butane respectively?

A 34.0 g mol^{-1}, C_2H_5, C_4H_{10}
B 29.0 g mol^{-1}, C_4H_{10}, C_2H_5
C 58.0 g mol^{-1}, C_2H_5, C_4H_{10}
D 12.0 g mol^{-1}, C_4H_{10}, C_2H_5

Question 7 (1 MARK)

Based on the following mass spectrum, identify the element that produces it by calculating the relative atomic mass.



A Selenium
B Arsenic
C Germanium
D Bromine

Question 8 (1 MARK)

How many moles of oxygen are there in 25.07 g of $CaCO_{3(s)}$?

A 0.2504 mol
B 1.567 mol
C 0.5223 mol
D 0.7513 mol

Question 9 (1 MARK)

How many moles of oxygen are there in 2.020×10^{23} molecules of $Ba(NO_3)_{2(aq)}$?

A 0.336 mol
B 0.671 mol
C 2.01 mol
D 5.37 mol

Question 10 (1 MARK)

Which of the following shows the balanced chemical equation between liquid glycerol ($C_3H_8O_3$) and oxygen gas (O_2) to produce carbon dioxide gas (CO_2) and water vapour (H_2O).

A $6CO_{2(g)} + 8H_2O_{(g)} \rightarrow 2C_3H_8O_{3(l)} + 7O_{2(g)}$
B $2C_3H_8O_{3(l)} + 7O_{2(g)} \rightarrow 6CO_{2(g)} + 8H_2O_{(g)}$
C $CO_{2(g)} + H_2O_{(g)} \rightarrow C_3H_8O_{3(l)} + O_{2(g)}$
D $C_3H_8O_{3(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (10 MARKS)

Through experimentation, it is determined that a neutrally charged atom of an element has an electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$.

a What is the name, chemical symbol, atomic number and relative atomic mass of this element? (2 MARKS)

b What block of the periodic table is this element found in? Justify your answer. (2 MARKS)

c Compare the electronegativity of this element with copper. (2 MARKS)

d Compare the metallic character of this element with zirconium. (2 MARKS)

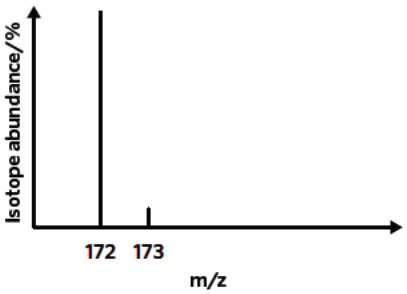
e Is this element more likely to react by gaining or losing electrons? (2 MARKS)

Question 12 (11 MARKS)

Esters are molecules, containing only carbon, hydrogen and oxygen, which are responsible for the smell of many fruits, including oranges. A confectionery company wanted to extract the ester from an orange in order to make orange smelling lollies.

a Combustion analysis of 3.07 g of this molecule produced 7.85 g of carbon dioxide and 3.21 g of water vapour. What is the empirical formula of the compound? (3 MARKS)

The following is a mass spectrum of the organic molecule. It reveals that the relative molecular mass of the compound is 172.0 g mol^{-1}.



b What is the molecular formula of this compound? (1 MARKS)

c Write a balanced chemical equation for the reaction of this molecule with oxygen to produce carbon dioxide and water vapour. (2 MARKS)

d The peak at m/z 173 is due to an isotope of an element. What is an isotope? (1 MARK)

e The relative abundance of the ^{13}C isotope, ^{2}H isotope and ^{17}O isotope is 1.1%, 0.02% and 0.04% respectively. Based on these values, explain why there is one small peak greater than the molecular mass of the compound. (2 MARKS)

f Food regulators have recently discovered nanoparticles in lollies. What are nanoparticles and why are concerns being raised? (2 MARKS)

Question 13 (9 MARKS)

There are two naturally occurring isotopes of copper, ^{63}Cu and ^{65}Cu. The relative atomic mass of copper is 63.6.

a In terms of abundance, which isotope is commonly found in nature? Justify your answer. (2 MARKS)

b Solid copper (I) oxide (Cu_2O) decomposes when heated into solid copper (Cu) and oxygen gas (O_2). Write the balanced chemical equation for this reaction. (2 MARKS)

c How many molecules of $O_{2(g)}$ will be produced from the decomposition of 1.00 mole of Cu_2O? (3 MARKS)

d What is the trend in first ionisation energy from copper to bromine? Explain this trend. (2 MARKS)

KEY SCIENCE SKILLS QUESTIONS (10 MARKS)

Question 14 (10 MARKS)

Antoine Lavoisier was a French chemist who conducted many experiments to prove the law of conservation of mass. One of these involved placing mercury in a glass jar and then heating it to see its reaction with oxygen.

a Write a possible aim for this experiment. (2 MARKS)

b Write a possible hypothesis for this experiment. (2 MARKS)

c Explain what data, and its type, would need to be collected in order to prove the law of conservation of mass. (3 MARKS)

d In one experiment, the jar wasn't sealed properly and so some of the product escaped. What effect would this have on the experiment and what type of error is it? (2 MARKS)

e Safety standards weren't very good in the 18th century. Suggest one precaution when handling toxic mercury. (1 MARK)

UNIT 1 AOS 1, CHAPTER 3

Metals

03

3A Metals

3B Variations of metallic substances

3C Reactivity of metals

Key knowledge

- the common properties of metals (lustre, malleability, ductility, heat and electrical conductivity) with reference to the nature of metallic bonding and the structure of metallic crystals, including limitations of representations; general differences between properties of main group and transition group metals
- experimental determination of the relative reactivity of metals with water, acids and oxygen
- the extraction of a selected metal from its ore/s including relevant environmental, economic and social issues associated with its extraction and use
- experimental modification of a selected metal related to the use of coatings or heat treatment or alloy production
- properties and uses of metallic nanomaterials and their different nanoforms including comparison with the properties of their corresponding bulk materials

Image: Zhao jiankang/Shutterstock.com

3A METALS

Metals have been commonly used in our daily lives including building bridges, towers and conducting electricity. In this lesson, we will learn about how metal particles are arranged in a solid and how they are extracted.

3A Metals | 3B Variations of metallic substances | 3C Reactivity of metals

Study design dot points

- the common properties of metals (lustre, malleability, ductility, heat and electrical conductivity) with reference to the nature of metallic bonding and the structure of metallic crystals, including limitations of representations; general differences between properties of main group and transition group metals
- the extraction of a selected metal from its ore/s including relevant environmental, economic and social issues associated with its extraction and use

Key knowledge units

Extraction of metals	1.1.7.1
Metallic bonding	1.1.5.1
Properties of metals	1.1.5.2

Key terms and definitions

Ore deposit in Earth's outermost layer containing metals and other minerals

Metallic bonding the electrostatic force of attraction between delocalised electrons and cations in a metallic lattice structure

Cation positive ion formed when an atom loses its valence electron(s)

Delocalised electrons electrons that freely move between metal cations in the metallic bonding model

Electrostatic force of attraction attractive force between charged particles

Malleability ability to deform under pressure without breaking

Ductility ability to be hammered or stretched into a thin shape without breaking

Heat conductivity ability to allow heat to pass through

Electrical conductivity ability to allow an electric current to flow through

Lustre shiny and glossy appearance

Lesson links

This lesson builds on:

- 2C The periodic table - part 1

Low ionisation energy of metals plays a role in the structure of solid metal.

Extraction of metals 1.1.7.1

OVERVIEW

Metals are extracted from deposits of minerals formed over long periods of time.

THEORY DETAILS

Most of the buildings, bridges like the Sydney Harbour bridge shown in figure 1 and towers all around the world have been built from metals.



Figure 1 Sydney Harbour bridge built from metals.

In fact, metals are the most common elements found in the periodic table as shown in figure 2.

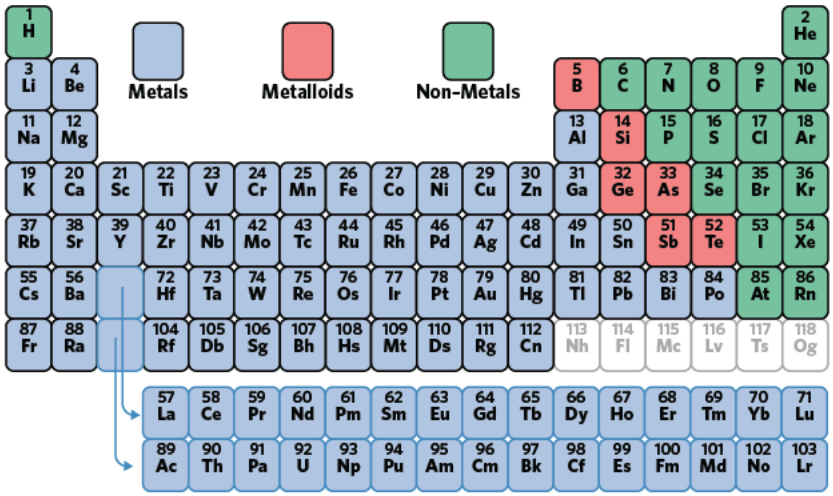


Figure 2 Metals in the periodic table.

Pure metals that we use are extracted from **ores** in the Earth's crust; the outermost layer of the Earth. Ores are deposits of minerals containing metals such as iron, zinc, aluminium, copper and gold that have formed over long periods of time. In ores, most metals except for platinum, silver and gold are present as compounds such as oxides, sulfides and silicates mixed with other minerals and rocky materials. For example, iron present in ore samples often exists as haematite (Fe_2O_3) and pyrite (FeS_2), and aluminium can be found in bauxite (Al_2O_3). Figure 3 shows a simplified process of extracting metals from their ores.

After being crushed to small pieces, ores are heated up to separate metal compounds from impurities. These metal compounds are then used to extract pure metal samples.

Different methods need to be used to chemically extract different metals from their compounds. In this lesson, we will only focus on carbon reduction, which is the method used to extract iron from its ores. Figure 4 shows the extraction of iron from iron oxide in the blast furnace.

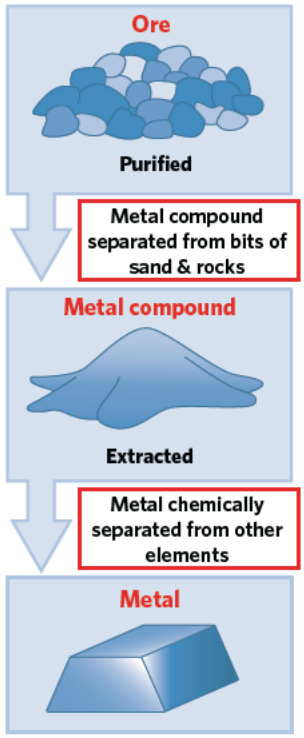


Figure 3 Overview of the extraction of a metal process.

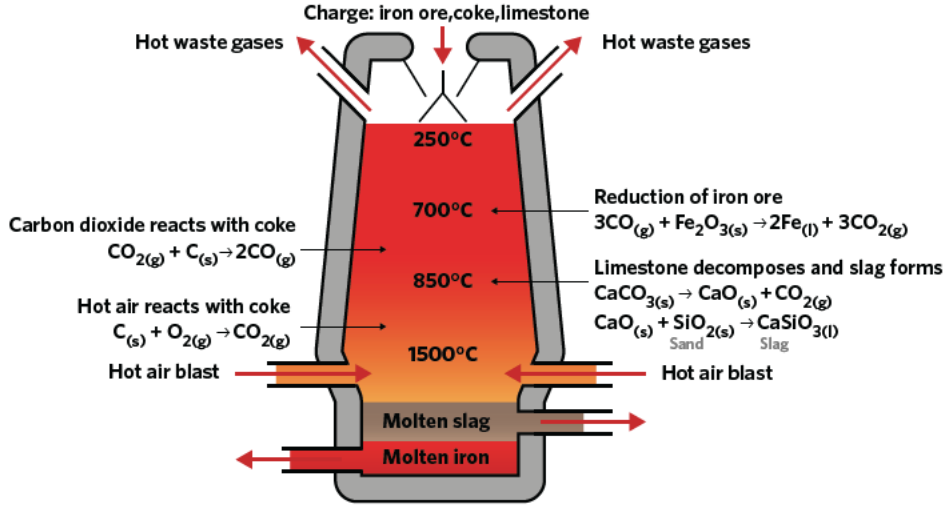


Figure 4 Extraction of iron from iron oxide in the blast furnace.

The extraction of iron requires raw materials such as coke, a solid mainly composed of carbon, and limestone, which is predominantly calcium carbonate. Iron ore, coke and limestone are added at the top of the furnace while hot air blasts are blown into the furnace near the base. The hot air reacts with coke to form carbon dioxide whilst the limestone breaks down into calcium oxide and carbon dioxide:

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

The carbon dioxide produced then reacts with coke to produce carbon monoxide:

$CO_{2(g)} + C_{(s)} \rightarrow 2CO_{(g)}$

After that, carbon monoxide reduces iron oxide Fe_2O_3 to Fe:

$3CO_{(g)} + Fe_2O_{3(s)} \rightarrow 2Fe_{(l)} + 3CO_{2(g)}$

The iron produced runs to the bottom of the furnace and ultimately melts into a liquid at approximately 1500°C. This can be summarised in the flowchart in figure 5.

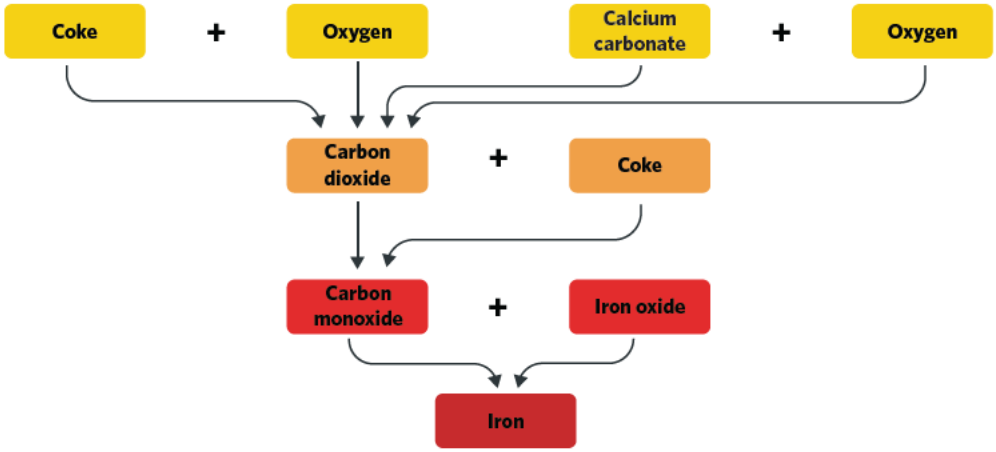


Figure 5 Flowchart of the metal extraction process.

Calcium oxide produced from the decomposition of limestone reacts with sand in the ores SiO_2 to form slag (calcium silicate):

$CaO_{(s)} + SiO_{2(s)} \rightarrow CaSiO_{3(l)}$

The slag floats at the top of the furnace and is drained off.

Even though metals are important materials used in a variety of fields, the extraction of metals has many negative impacts on the environment. The large amount of rocks and sand collected to extract metals leads to the destruction of the landscape which then causes flooding, soil erosion and water pollution. Furthermore, gases released during the extraction of metals such as carbon dioxide and sulfide oxide contribute to global warming and acid rain, while the dust released decreases the air quality.

The extraction and exportation of minerals has played an important role in the economic growth of Australia. However, the whole process is also quite costly and therefore also needs to be taken into consideration.

Regarding the social impacts, even though the mining industry has created a lot of jobs for communities living near mining sites, people living there are affected by noise, air, land and water pollution. There are also adverse health effects such as respiratory and digestive diseases and most seriously cancer that could arise for those living close to these mining sites. An example of this is the communities living close to gold mining areas in Brazil, where it has been shown that locals experience higher risks of cancer due to thousands of tonnes of mercury released into the environment.

Metallic bonding 1.1.5.1

OVERVIEW

Metallic bonding model explains how metal atoms bond to each other when in solid form.

THEORY DETAILS

As we learned in lesson 2C, metals have low ionisation energies and therefore require a relatively small amount of energy to remove valence electrons from their atoms. As a result, it is easy for metal atoms to lose electrons from their valence shell and become positive ions called cations. Figure 6 shows how a metal atom loses its electrons.

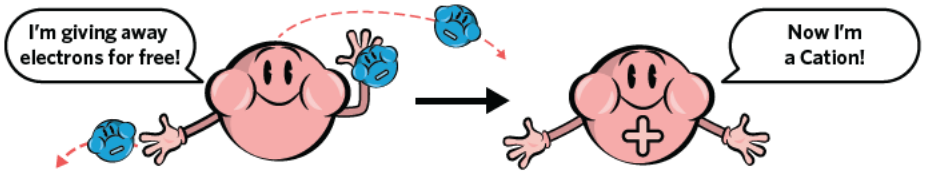
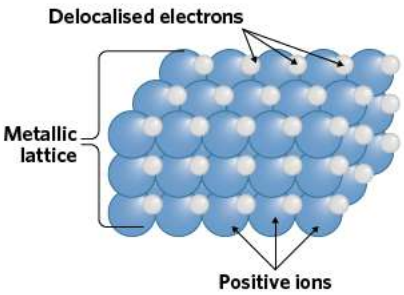


Figure 6 Illustration of a metal atom losing electrons.

When in solid form, metals are organised in a lattice structure as shown in figure 7. When organised in this way, electrons removed from the valence shells of metal atoms are able to move freely within the lattice structure between the positively charged metal cations. These electrons are called **delocalised electrons**. The **electrostatic force of attraction** between the negatively charged delocalised electrons and positively charged metal cations holds the metal cations closely to each other in fixed positions and maintains the structure of the metal. The electrostatic force of attraction between the delocalised electrons and cations is referred to as **metallic bonding**.



Positive ions form the crystal lattice, while delocalised electrons move freely throughout the lattice, gluing the whole lattice together.

Figure 7 Metallic bonding model.

Properties of metals 1.1.5.2

OVERVIEW

Metals exhibit different properties that allow them to be used for many different purposes.

THEORY DETAILS

As the structure of metals can be explained using the metallic bonding model, most metals share common properties as shown in table 1.

Table 1 Properties of metals.

Property	Description
Malleable	Can be bent and hammered into different shapes without breaking
Ductile	Can be stretched into thin wire
Heat conductivity	Allows heat to pass through
Electrical conductivity	Allows electric current to flow through
High melting point	High melting temperatures
Lustrous	Looks shiny and reflective

The ability of metals to be ductile and malleable is due to the electrostatic forces of attraction between the cations and delocalised electrons. When a mechanical force is applied on metals to bend, hammer or to change their shapes, layers of cations are able to slide past each other whilst still being held in place by the electrostatic force of attraction with the delocalised electrons as shown in figure 8.

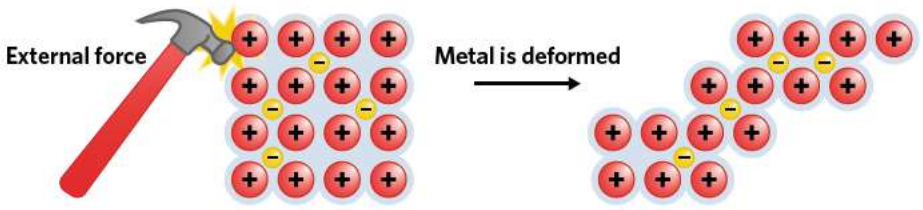


Figure 8 The effect of an applied force on the metallic bonding structure.

As a result, the structure can take another shape without breaking.

Metals are considered to be good heat conductors, which can be explained through the metallic bonding model. When a piece of metal absorbs heat, the rise in temperature causes the increase in kinetic energy of delocalised electrons and metal ions. With this increase in energy, the metallic ions can vibrate more rapidly and delocalised electrons can move faster, transferring the kinetic energy to nearby ions and electrons, creating a domino effect that allows the energy to be passed through the metal as seen in figure 9.

When the piece of metal is in contact with another object, the kinetic energy carried by moving electrons will transfer to the object in the form of heat. This is why we feel hot when touching a heated saucepan.

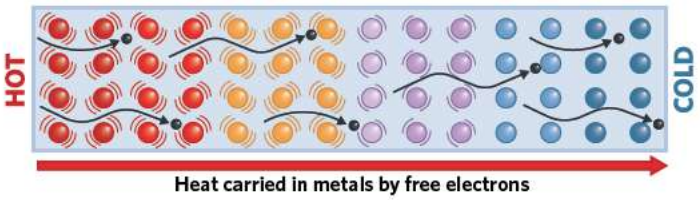


Figure 9 The conduction of heat through a metallic substance.

As we can appreciate, the conduction of heat requires particles to transfer the energy from a hot region to a cold region relatively quickly. Therefore, if we were to place a piece of metal into a beaker of iced water, the metal would be able to quickly transfer heat energy from the metal (hot region) to the water (cold region). As a result, the metal would feel cold relatively quickly.

The metallic bonding model can also be used to explain the electrical conductivity of metals, which refers to a metal's ability to conduct electricity. As shown in figure 10, in a basic circuit, electrons are continuously moving away from the negative electrode of a power source towards the positive electrode. Due to the fact that metals contain delocalised electrons when connected to the circuit, there is a movement of delocalised electrons away from the negative electrode and towards the positive electrode, which generates a current in the circuit. In the case of the circuit in figure 10, this results in the light being turned on. This property is why electrical wires are made from metallic substances.

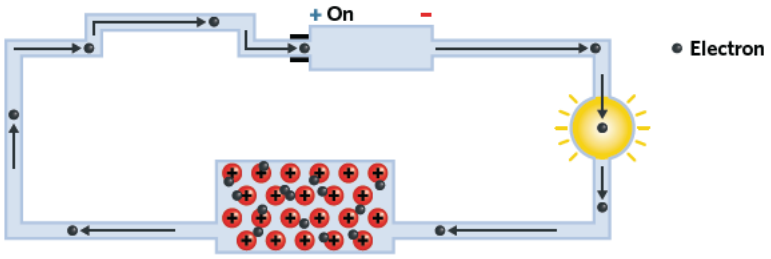


Figure 10 Using the metallic bonding model to explain the electrical conductivity of metals.

Metals are generally hard and have relatively high melting points. This is due to the strong electrostatic force of attraction that holds metal cations and delocalised electrons closely and tightly together. As a result, a large amount of heat is required in order to break the metallic bonds and deform the structure of metal lattices.

The shiny or lustrous nature of metals is due to the delocalised electrons, which are able to reflect the light of any wavelength.

Worked example 1

Use the metallic bonding model to explain why copper is widely used make wires that conduct electricity.

What information is presented in the question?

Use the metallic bonding model.

Copper is widely used to make wires that conduct electricity.

What is the question asking us to do?

Explain why copper wire is widely used to make wires that conduct electricity.

What strategy(ies) do we need in order to answer the question?

1. Describe the metallic bonding model of copper.
2. Use the metallic bonding model to describe what occurs when copper wire is connected to an electrical source.
3. Conclude the electrical conductivity of copper.

Answer

In the metallic bonding model of copper, delocalised electrons move freely in between copper cations. Therefore, when one end of the copper wire is connected to the positive electrode and the other end is connected to the negative electrode of an electrical source, delocalised electrons move away from the negative electrode and towards the positive electrode, which generates a current in the circuit that lightens the bulb. Hence, copper is widely used to make wires that conduct electricity.

Differences between main group metals and transitions metals

Transition metals are the elements located between groups 2 and 13 in the periodic table as shown in figure 11.

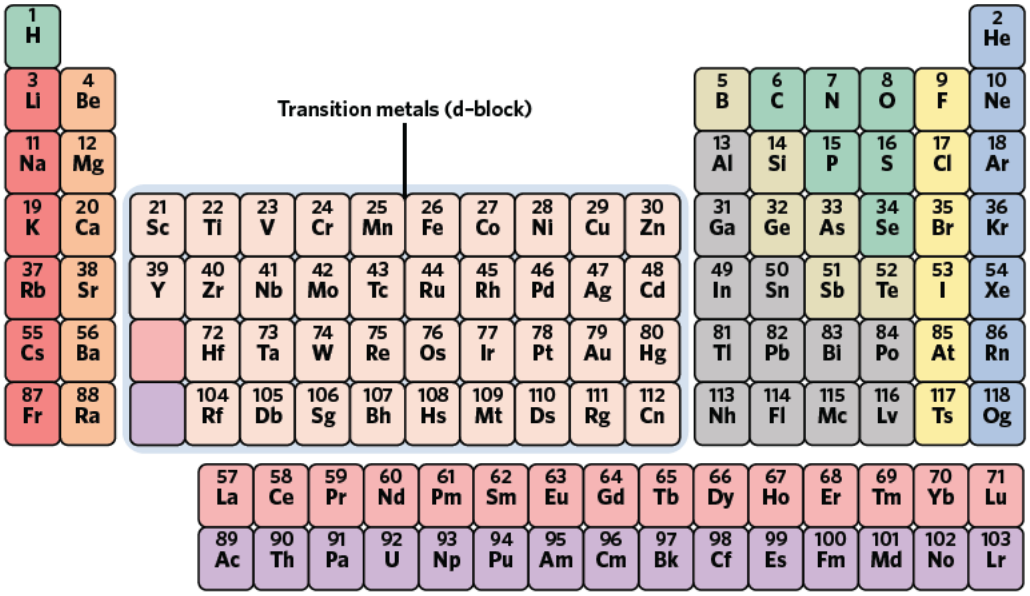


Figure 11 Locating transition metals in the periodic table.

Even though transition metals exhibit lustre, malleability, ductility, heat and electrical conductivity, they exhibit properties that are different from main group metals. Compared to group 1 and group 2 metals, transition metals have small atomic sizes due to their larger core charges. This means that there is less space between the cations of transition metals in the metallic bonding model and therefore they are packed more tightly and closely together shown in figure 12. As a result, relative to group 1 and group 2 metals transition metals:

- have higher densities.
- are harder.
- have higher melting points because it requires more heat to break their tightly packed lattices.

Some transition metals also have strong magnetic properties.

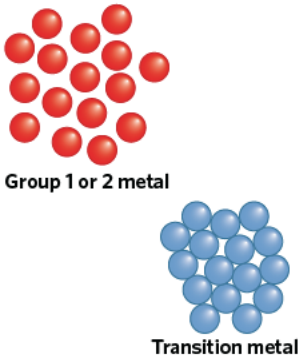


Figure 12 Packing of transition metal atoms compared to group 1 and group 2 metal atoms.

Limitations of the metallic bonding model

Even though the metallic bonding model can be used to explain some properties of metals, it isn't able to explain the variations in the properties shown by different metals. Particularly, it cannot be used to explain:

- the different melting points of different metals.
- the difference in electrical conductivity of different metals.
- magnetic properties shown by a select number of metals.

As the explanation for these properties falls outside the study design, we will not be going into any further detail about the metallic bonding model.

Theory summary

- In the extraction of metals, ores are first crushed into small pieces and heated up to separate metal compounds from impurities and finally pure metals are chemically extracted from metal compounds like oxides and sulfides.
- Iron is extracted from iron oxides in the blast furnace.
- The extraction of metals can have many impacts on the environment, the economy and the society.
- In the metallic bonding model, the electrostatic force of attraction between delocalised electrons and metal cations keep metal cations closely packed together.
- Metals exhibit lustre, malleability, ductility, heat and electrical conductivity which can be explained by the metallic bonding model.

3A QUESTIONS

Theory review questions

Question 1

Metals are extracted from samples similar to the image provided.



Image: Kletr/Shutterstock.com

These samples are

A rocks and sands formed over a long period of time in the Earth's crust.

B deposits of minerals formed over a long period of time in the Earth's crust.

Question 2

The correct order of the metals extraction process is:

A Ores are crushed into small pieces, ores are heated up to separate metal compounds from sands and rocks, pure metals are chemically extracted from metal compounds.

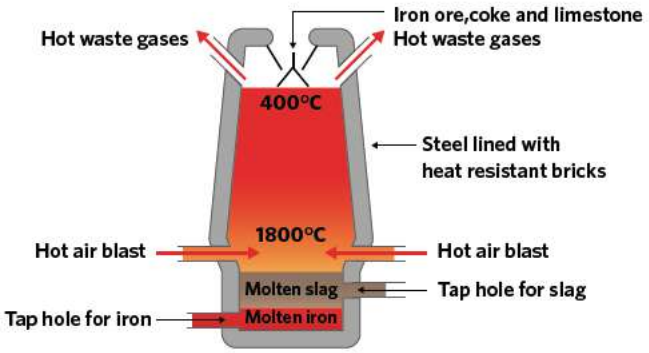
B Ores are heated up to separate metal compounds from sands and rocks, ores are crushed into small pieces, pure metals are chemically extracted from metal compounds.

Question 3

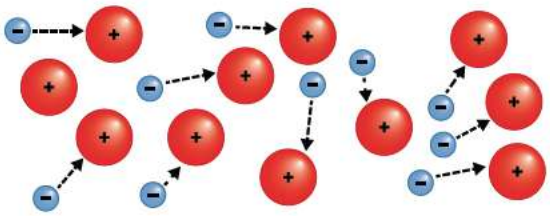
Iron is chemically extracted from its oxide in the blast furnace. In the blast furnace, hot air reacts with coke to form CO_2 which then reacts with coke to form CO. What is the role of CO in the extraction of iron?

A It maintains the cycle of hot air in the blast furnace.

B It reduces iron oxide to iron.



Use the following information to answer questions 4–5.



Question 4

The small freely moving particles illustrated in the image provided are referred to as

A localised electrons.

B delocalised electrons.

Question 5

In the metallic bonding model, metal cations are held in fixed positions and kept closely packed together by

A the electrostatic force between localised electrons and metal cations.

B the electrostatic force between delocalised electrons and metal cations.

Question 6

Some properties of metals can be explained by the metallic bonding model.

A True

B False

Question 7

Metals are malleable due to

A the electrostatic force between delocalised electrons and metal cations holds metal cations closely and tightly together but allows cations to slide over each other.

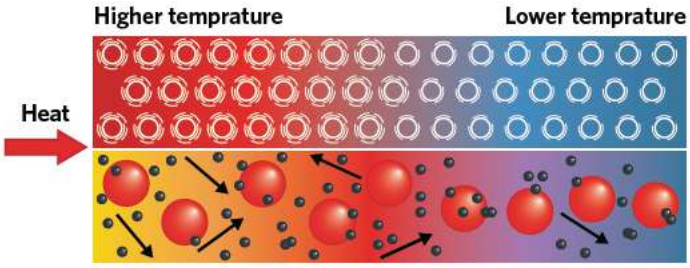
B the electrostatic force between localised electrons and metal anions holds metal cations closely and tightly together but allows cations to slide over each other.

Question 8

When a piece of metal is heated, the _____ electrons absorb the heat to increase their _____. Therefore, they move faster and transfer the kinetic energy to the _____ electrons nearby. As a result, heat is transferred along the piece of metal.

A delocalised, kinetic energy, delocalised

B localised, chemical energy, delocalised

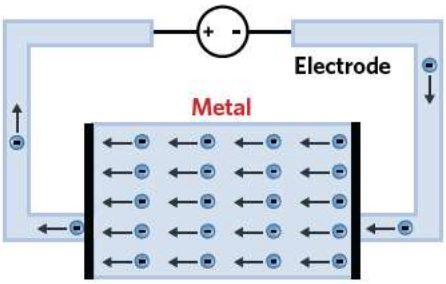


Question 9

The electrical conductivity of metals can be explained by the movement of

A delocalised electrons away from the positive electrode and towards the negative electrode.

B delocalised electrons away from the negative electrode and towards the positive electrode.



Question 10

Some metals have a melting point that is over 1000°C. This suggests that

A metallic bonds are relatively strong.

B delocalised electrons are able to move quickly through the lattice.

Deconstructed questions

As shown in the periodic table, there are many different types of metals that exist on earth. The table provided shows the melting points of different metals

Metal	Melting point °C
Iron	1535
Aluminium	660
Copper	1083
Magnesium	650

Question 11 (1 MARK)

Based on the information provided by the table, which metal could be considered as being most resistant to heat?

A Iron

B Copper

C Aluminium

D Magnesium

Question 12 (1 MARK)

With regards to the metallic bonding, order each metal based on the strength of their bonds from strongest to weakest.

A Iron, aluminium, copper, magnesium

B Magnesium, copper, aluminium, iron

C Iron, copper, aluminium, magnesium

D Copper, iron, magnesium, aluminium

Question 13 (3 MARKS)

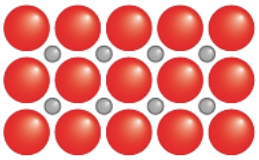
A building company wanted to build the basic frame structure for a four storey concrete apartment block. Which of the metals listed above would allow for the most sturdy structure?

Exam-style questions

Within lesson

Question 14 (4 MARKS)

The structure of metals allow them to exhibit a wide range of properties.



a Label the structure shown assuming that it is a representation of the structure of a metal. (2 MARKS)

b Gold is known to be very malleable. As a result, it is able to be made into many different types of products. Explain, using the model shown, how gold is able to be quite malleable. (2 MARKS)

Question 15 (4 MARKS)

Describe the impacts of the extraction of metals on

a the environment. (2 MARKS)

b the communities residing near the extraction sites. (2 MARKS)

Question 16 (3 MARKS)

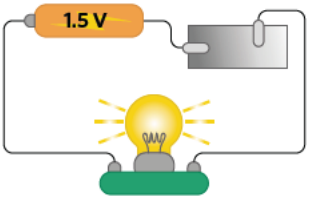
Nick conducted an experiment to investigate the properties of metals. He placed a metal spoon into a bowl of boiled water. After 30 seconds, he had a 'burning' feeling when he touched the spoon.

a What can Nick conclude about the property of metals? (1 MARK)

b Explain the identified property of metals using the metallic bonding model. (2 MARKS)

Question 17 (3 MARKS)

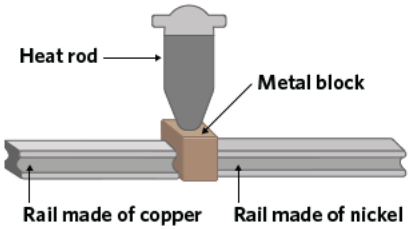
A student argued that metals conduct electricity because metal cations can move from the positive electrode to the negative electrode. Is the student's explanation correct? Justify your answer with the metallic bonding model.



Question 18 (5 MARKS)

With a rising population, the City of Dalton commissioned an engineering company to develop a railway using materials that was able to withstand high levels of heat whilst being a low conductor of heat.

As part of their investigation, they developed an experiment involving two metals as shown in the diagram.



Two pieces of rail developed using different metals were joined to a metal block. The engineers then applied heat to the metal block through an electric heat rod at increasing temperatures. Observations made during the experiment are shown in the table.

Temperature (°C)	Copper rail	Lead rail
200	unchanged	unchanged
400	unchanged	the rail was beginning to lose its shape
600	It was glowing red but structure remained unchanged.	No solid metal was left, there was a lot of hot liquid in place of where the piece of rail used to be.

a Which of the two metals used would be considered as having a higher melting point? Explain. (3 MARKS)

b Is it possible to use the results provided to compare the heat conductivity of both metals? Explain. (2 MARKS)

Multiple lessons

Question 19 (4 MARKS)

A biomedical engineer wants to choose a metal to design hip implants. Due to the wide range of movement of the human hips, the hip implants are required to withstand strong bending force without breaking. The engineer is considering choosing either magnesium or titanium.

a Should the engineer choose magnesium or titanium to create hip implants? (1 MARK)

b Justify your answer. (3 MARKS)

Key science skills

Question 20 (4 MARKS)

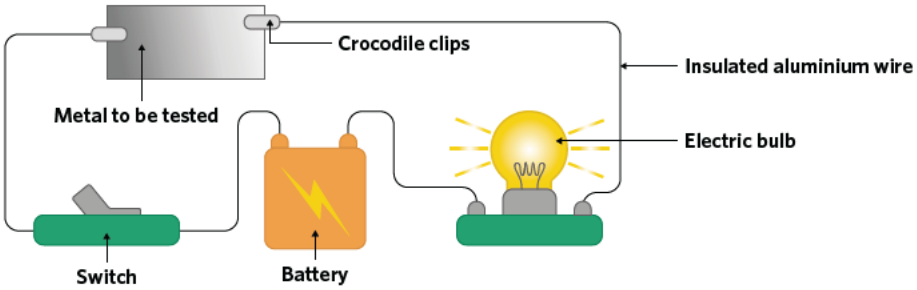
Gabriella conducted an experiment to confirm if her ring is entirely made of silver as the jeweller said.

Materials

- 1 × Electric bulb
- 1 × Battery
- 1 × Switch
- Insulated aluminium wires
- 2 × Crocodile clips
- The ring
- 1 × Piece of pure silver

Method

The diagram of the electrical circuit designed by Gabriella:



Results

Object	Observation
The ring	Dull light from the bulb
Piece of pure silver	Strong and dazzling light from the bulb

a From the obtained results, what can Gabriella conclude about the material of her ring? Explain your answer. (3 MARKS)

b Gabriella's friend told her that her conclusion might not be correct because the electricity might not be completely transferred through the wires. Suggest one change Gabriella can make to increase the accuracy of her experiment. (1 MARK)

Exam-style question hints

14a: In general, metals have low ionisation energies. **14b:** The nature of bonds allows substances to exhibit particular properties. **15a:** The required materials and products of extraction of metals can impact the environment. **15b:** Impacts on the environment can lead to health issues. **16a:** 'Burning' feeling when touching the spoon. **16b:** Increased temperature can increase kinetic energy of delocalised electrons and metals in the metallic bonding model. **17:** Delocalised electrons can move freely in the metallic bonding model. **18a:** Changes to the structure of a substance can indicate its properties. **18b:** Even with similar structures, metals exhibit slightly different properties. **19b:** Atomic sizes can affect how closely metal atoms are packed together and therefore affect how much force metals can withstand. **20a:** Same observations should be obtained from the same pure metals. **20b:** Different metal wires have different electrical conductivities.

3B VARIATIONS OF METALLIC SUBSTANCES

In the previous lesson, we explored the properties of pure metals. In this lesson, we will learn about the effects of modifying metals, different forms of metallic nanomaterials as well as their properties and applications.

3A Metals | 3B Variations of metallic substances | 3C Reactivity of metals

Study design dot points

- experimental modification of a selected metal related to the use of coatings or heat treatment or alloy production
- properties and uses of metallic nanomaterials and their different nanoforms including comparison with the properties of their corresponding bulk materials

Key knowledge units

Modifying metals	1.1.8.1
Different forms of metallic nanomaterials	1.1.9.1
Properties and uses of metallic nanomaterials	1.1.9.2

Key terms and definitions

Alloy material that is made by combining a metal with another chemical element which can be a metal or a nonmetal

Alloying agent element combined with a main metal to make an alloy

Interstitial alloy alloy in which alloying agent atoms stay in the gaps between main metal cations

Substitutional alloy alloy in which alloying agent atoms replace some main metal cations

Annealing heat treatment in which metals are heated and left to cool down slowly

Quenching heat treatment in which metals are heated and left to cool down quickly

Tempering heat treatment in which metals are heated and left to cool down in the air

Lesson links

This lesson builds on:

- 3A Metals

Modification of metals can change the structure of metal lattices and therefore change their properties.

Modifying metals 1.1.8.1

OVERVIEW

The structures and properties of metals can be changed by modifying metals with different methods.

THEORY DETAILS

Alloys

An **alloy** is a material that is made by combining a metal with another chemical element, which can be either a metal or a nonmetal. A typical alloy consists of at least 90% of one main metal and a small proportion of another element called an **alloying agent**. Most alloys are produced by the following steps:

- Melt the substances.
- Mix the collected liquids together.
- Cool the mixture.

There are two main types of alloys, substitutional alloys and interstitial alloys, which are categorised based on their structures.

An **interstitial alloy** is formed when alloying agent atoms fill in the gaps between main metal atoms. For this to be achieved, alloying agent atoms need to be much smaller than the main metals atoms. The structure of an interstitial alloy is shown in figure 1 where the main metal cations are shown in blue and the alloying agent atoms are shown in pink.

One of the most common interstitial alloys is carbon steel, which is widely used in the construction of railways, roads, bridges and buildings. It is made up of iron and a small amount of carbon.

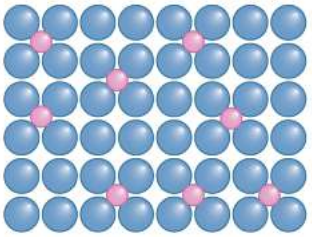


Figure 1 The structure of an interstitial alloy.

Carbon steel has some advanced properties compared to pure iron. In pure iron, iron cations are arranged in regular layers that can slide over each other easily. Therefore, it is easy for pure iron to be deformed by applying force. Contrastingly, in carbon steel, carbon atoms which are smaller than iron cations distort regular layers of the iron cations, meaning that it requires more force to make iron cations slide over each other. As a result, carbon steel is harder, stronger, less malleable and less ductile than pure iron.

As the proportion of carbon within the structure increases, the hardness of carbon steel will also increase while the malleability and ductility will decrease, as shown in table 1.

Table 1 Comparison of the different types of carbon steel.

Type of carbon steel	Proportion of carbon	Properties
Low carbon steel	Less than 0.3%	Hard, easily deformed and shaped
Medium carbon steel	0.3–0.45%	Harder, stronger and less ductile
High carbon steel	0.45%–0.75%	Very hard, strong and more brittle
Very high carbon steel	Up to 2.5%	Very hard, strong and brittle

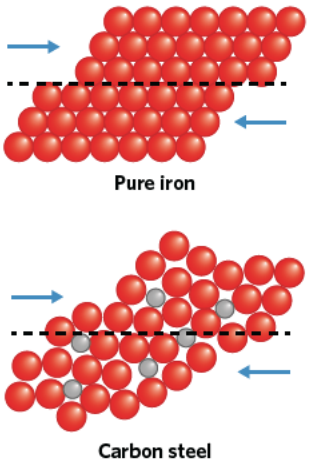


Figure 2 Demonstration of the deformation of pure iron and carbon steel.

In a **substitutional alloy**, some main metal cations are replaced by atoms of similar size. This is shown in figure 3, where the main metal cations are shown in blue and the alloying agent atoms are shown in pink.

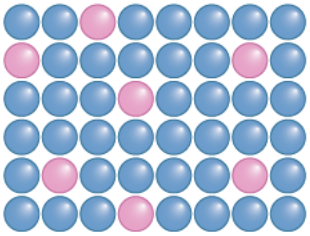


Figure 3 The structure of a substitutional alloy.

Stainless steel, which is made of iron and a small amount of nickel or chromium is an example of substitutional alloy. Although the atoms in stainless steel are of similar size, the slight size difference between the iron, nickel or chromium cations results in a structure that is stronger and harder than pure iron. Therefore, regular layers of iron cations are slightly distorted in the alloy lattice, meaning that it requires more force to make all these metal cations slide over each other in the alloy lattice compared to a pure iron lattice. The increased strength and hardness of stainless steel is also due to the fact that nickel or chromium cations added are strongly attracted to delocalised electrons within the lattice.

Heat treatment

Applying heat to metals can also lead to a change in the structures and properties of metals. There are three common methods to modify metals by heat treatment, **annealing**, **quenching** and **tempering**, as highlighted in table 2.

Table 2 Heat treatment methods.

Treatment	Process	Change in structure	Change in properties
Annealing	Heating a metal to a specified temperature where recrystallisation occurs and allowing it to cool down slowly to room temperature	Formation of large metal crystals	Increased softness and ductility
Quenching	Heating a metal to a specified temperature where recrystallisation occurs and quickly cooling it down in water or oil bath	Formation of small metal crystals	Increased hardness and strength
Tempering	Heating a quenched metal to below recrystallisation temperature and allowing it to cool down in the air	Formation of medium metal crystals	Increased hardness, brittleness, strength and ductility

When metals are heated up to a specified temperature, metal crystals combine with each other. When they are cooled down, metal crystals are reformed. As seen in table 2, the more slowly a metal is cooled down, the bigger the crystals are formed as they are given more time to grow in size. It is also important to note that larger crystals are more flexible and therefore are easier to shape, while smaller crystals are stiffer and less easy to shape. As a result, annealing increases softness and ductility while quenching increases hardness, strength and brittleness of metals.

Coatings

Metals can react with a variety of substances such as oxygen, acid and water, which makes the surface of metals undergo rusting or corrosion. This will be discussed in greater depth in lesson 3C. To protect the surface of a metal from being damaged over time by reactive substances, a protective layer is coated on the surface of the metal. This extra layer is called a metal coating and is commonly made from aluminium, nickel, zinc, chromium or cadmium. Table 3 shows some widely used coatings and their uses.

Table 3 Types of metal coating.

Types of coating method	Process of coating	Uses
Anodising	Adding a thick metal oxide layer to the surface of a metal	More resistant to corrosion
Galvanising	Applying a layer of zinc to the surface of a metal	More resistant to corrosion
Electroplating	Using an electrical current to adhere a solution of cadmium and chromium to a metal	More resistant to corrosion and better appearance
Powder coating	Applying a dry powder to the surface of a metal	Resistant to corrosion, smooth and pleasing surface

Different forms of metallic nanomaterials 1.1.9.1

OVERVIEW

Pure metals can be arranged to produce nano-sized structures called metallic nanomaterials.

THEORY DETAILS

Not only can metals be seen in massive structures such as towers, bridges and buildings, as shown in lesson 3A, but they can also be used to create extremely tiny substances which are shorter or smaller than one millionth of a meter. These substances are called metallic nanomaterials and they can be viewed under a microscope. There are three main types of metallic nanomaterials including metallic nanoparticles, metallic nanorods and metallic nanowires.

Metallic nanoparticles

Metallic nanoparticles are spherical particles which are made of metals and have a size range of 1 to 100 nanometers (nm), where 1 nm = 10^{-9} m. They are commonly used to deliver drugs, kill bacteria and remove pollutants in water.

Metallic nanorods

Metallic nanorods are rod-like shaped structures as seen in figure 4 whose sizes are from 1 to 100 nm. They are utilised in photothermal therapy in cancer treatment.

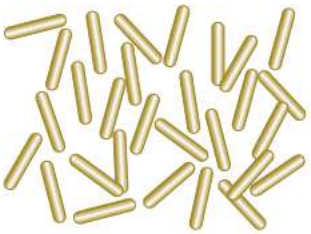


Figure 4 Gold nanorods.

Metallic nanowires

Metallic nanowires are wire-shaped structures which have nano-sized diameters but can be very long. Figure 5 demonstrates the appearance of silver nanowires under a microscope. One use of metallic nanowires is to design electrical systems in energy-efficient buildings.

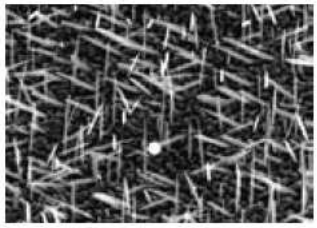


Figure 5 Silver nanowires.

Properties and uses of metallic nanomaterials 1.1.9.2

OVERVIEW

Metallic nanomaterials possess special properties that enable them to be applied in many fields.

THEORY DETAILS

Metallic nanoparticles

Metallic nanoparticles with spherical shapes exhibit some advanced properties in comparison to bulk metals. Due to their tiny size, there is a much larger number of nanoparticles that can fit within a space with a particular volume. Therefore, they have a higher surface area to volume ratio than bulk metals, meaning that they have a higher surface area than bulk metals of similar or same volume.

Because nanoparticles are tiny, there is a limited number of metal cations in the structure of metallic nanoparticles, which also means that there is a relatively small amount of delocalised electrons present between the metal cations. Consequently, nanoparticles have properties in between those of metals and non-metals including unique electrical, magnetic, thermal and mechanical properties, however these details are beyond the scope of this course.

Due to the special arrangement of nanoscale particles, metallic nanoparticles have different optical properties from bulk metals which allow them to absorb a wide range of visible ultraviolet (UV) light that cannot be absorbed by bulk metals.

The high surface area to volume ratio enables metallic nanoparticles to be used in drug delivery, as shown in figure 6. For example, gold nanoparticles are used to deliver chemotherapeutic drugs used in cancer treatment. Furthermore, the use of gold nanoparticles also works to hide the drug being introduced from the body's immune system, ensuring that it will not be broken down before it carries out its function.

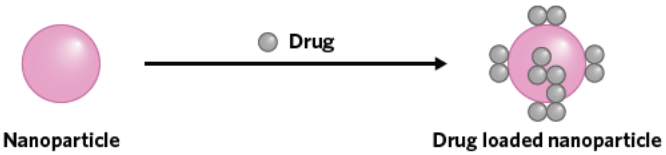


Figure 6 Example of the use of a gold nanoparticle in drug delivery.

Due to the high surface area to volume ratio of nanoparticles, they can diffuse at a fast rate even at low temperature. Therefore, they are placed on the surface of drugs with low solubility to enable these drug particles to move faster through the bloodstream to reach target cells more quickly. Moreover, increasing the diffusion rate of drugs by using metallic nanoparticles without increasing the temperature can protect healthy cells from being damaged.

The high surface area to volume ratio also enables metallic nanoparticles to readily diffuse through the membranes of bacteria, subsequently disrupting their cell membranes and other structures and ultimately killing them. Because of this, some metallic nanoparticles, such as silver nanoparticles, have been embedded in wound dressings to kill bacteria at sites of injury, which can speed up the recovery process, as shown in figure 7.

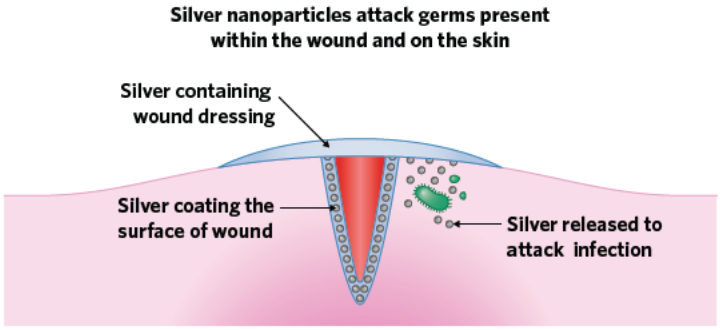


Figure 7 Role of silver nanoparticles in wound dressings.

Iron nanoparticles have been utilised to remove pollutants and contaminants in water and soil. Carbon tetrachloride, a molecule present in cleaning and degreasing fluids, is a toxic chemical which can contaminate in water and solid which can cause cancer in humans and animals. Iron nanoparticles can quickly and effectively break down carbon tetrachloride in water and soil, therefore improving water and soil quality.

Metallic nanorods

Metallic nanorods with rod-like shapes, especially gold nanorods have unique optical characteristics enabling a wide range of light absorption. This property makes them useful in the field of biomedical imaging, where their optical properties aid with diagnosing cancer.

Gold nanorods have also been used for photothermal therapy in cancer treatment. When provided with infrared light, they can release a relatively large amount of heat in the human body that can locally destroy cancerous cells without damaging healthy cells.

Metallic nanowires

Metallic nanowires with long wire-like shapes are excellent conductors of electricity. Because of this, they have been popular materials used to produce electrodes, which are used as electrical conductors in various contexts. Moreover, due to their good heat conductivity, metallic nanowires have been widely used to design energy-efficient buildings. They can effectively absorb, store and release heat and therefore can replace air-conditioning systems that consume a large amount of electricity.

Theory summary

- Metals can be modified to make interstitial alloys or substitutional alloys.
- There are three main methods to modify metals using heat including annealing, quenching and tempering.
- Coatings can be applied to the surface of metals to slow down or prevent corrosion.
- There are three main types of metallic nanomaterials including metallic nanoparticles, metallic nanorods and metallic nanowires, all of which have unique and advanced properties that enable their applications in various fields.

3B QUESTIONS

Theory review questions

Question 1

Metals can be modified to make alloys.

A True

B False

Question 2

A substitutional alloy can be produced from zinc by

A adding some atoms of a different metal that have a much smaller size than zinc atoms.

B replacing some zinc atoms with some atoms of a different metal that have a similar size to zinc atoms.

Question 3

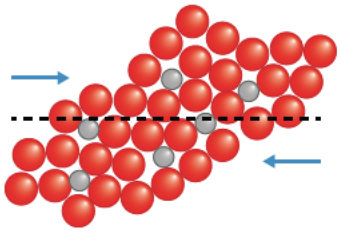
In an interstitial alloy, atoms of an alloying agent that have _____ size than main metal atoms sit in between main metal atoms.

A larger

B smaller

Question 4

An interstitial alloy is harder and less malleable than its pure metal because _____ alloying agent atoms distort the regular layers of main metal cations, meaning that it requires _____ to make main metal cations to slide over each other.



A small, more force

B large, less force

Question 5

The _____ time metals are left to cool down after being heated to a particular temperature, the _____ the crystals can be formed.

A less, bigger

B more, bigger

Question 6

Why does annealing create much larger metal crystals than quenching?

A Because a heated metal is allowed to cool down slower in annealing than in quenching.

B Because a heated metal is allowed to cool down in a higher pressure in annealing than in quenching.

Question 7

What is the main purpose of adding a coating on the surface of a metal?

A To increase heat conductivity.

B To prevent or slow down the corrosion.

Question 8

Electroplating involves the process of

A applying a dry powder to the surface of a metal.

B using an electrical current to adhere ions from a solution of cadmium and chromium to a metal.

Question 9

Metallic nanoparticles have properties in between those of metals and non-metals because

A there is a small number of delocalised electrons between metal cations due to a limited number of metal cations.

B there is a small number of delocalised electrons between metal cations due to a large number of metal cations.

Question 10

Metallic nanoparticles are highly applicable in drug delivery because they can quickly diffuse through cell membranes. Which characteristic enables them to readily diffuse through cell membranes?

A High surface area to volume ratio.

B Broad range of light absorption.

Deconstructed questions

Use the following information to answer questions 11–13

A structural engineer is given a task of choosing a material to make construction wires which can easily be bended and reshaped. The engineer is considering using one of the carbon steel types shown in the table.

Name of carbon steel	Percentage of carbon
AISI 1010	0.1%
AISI 1040	0.37–0.44%
AISI 1080	0.75–0.88%
AISI 1095	0.90–1.04%

Question 11 (1 MARK)

Which of the listed carbon steel types has the lowest percentage of carbon?

A AISI 1010

B AISI 1040

C AISI 1080

D AISI 1095

Question 12 (1 MARK)

Which of the following statements is correct about carbon steel?

A Added carbon atoms make it easier for metal cations to slide over each other.

B Added carbon atoms make it harder for metal cations to slide over each other.

C The fewer carbon atoms added, the less force required to deform and reshape carbon steel.

D Both B and C are correct.

Question 13 (3 MARKS)

Which type of listed carbon steel should the engineer choose? Explain your answer with reference to the lattice structure of carbon steel.

Exam-style questions

Within lesson

Question 14 (6 MARKS)

Stainless steel is made from iron and a small amount of nickel. It is commonly used to make cutlery including knives, forks and spoons.

a What type of alloy is stainless steel? Justify your answer. (2 MARKS)

b Describe two reasons why stainless steel is stronger and harder than pure iron. (4 MARKS)

Question 15 (4 MARKS)

Steve wants to make his own bicycle frame from a massive piece of iron by following 2 steps:

1 Make the piece of iron more malleable to shape it into a bicycle frame.

2 Make sure that the frame can withstand a heavy load.

a What heat treatment should Steve do in step 1? Explain your answer. (2 MARKS)

b What heat treatment should Steve do in step 2? Explain your answer. (2 MARKS)

Question 16 (5 MARKS)

There are three main types of metallic nanomaterials: metallic nanoparticles, metallic nanorods and metallic nanowires.

a Describe the one key similarity and difference in the structures of these three types of metallic nanomaterials. (3 MARKS)

b Which property enables metallic nanowires to be widely used to design energy-efficient buildings? Explain your answer. (2 MARKS)

Question 17 (4 MARKS)

Krissy, a biomedical scientist, conducted an experiment to compare the drug delivery capabilities of three different types of metallic nanoparticles. She compared their surface area to volume ratios and the temperatures required to activate them. The obtained results are shown in the given table.

Type of metallic nanoparticles	Surface area to volume ratio	Temperature
A	110 : 1	37°C (normal body temperature)
B	147 : 1	37°C (normal body temperature)
C	148 : 1	39°C

According to the obtained results, which type of metallic nanoparticles is the best for drug delivery? Justify your answer by explaining why one type is better than the other two.

Multiple lessons

Question 18 (4 MARKS)

A sample of carbon steel which is composed of 90% iron and 10% carbon by mass has a total mass of 175 g.

a Calculate the mass and the number of moles of iron present in the sample. (2 MARKS)

b Calculate the mass and the number of moles of carbon present in the sample. (2 MARKS)

Question 19 (6 MARKS)

Stainless steel is commonly made by combining iron with a small amount of nickel.

a Write the electron configurations of iron and nickel atoms. (2 MARKS)

b Which of the two atoms has a larger radius? Justify your answer. (2 MARKS)

c Which of the two atoms is more electronegative? Explain your answer. (2 MARKS)

Key science skills

Question 20 (6 MARKS)

After having studied about the tempering of metals, Warren realised that the tempering technique is also used to change the properties of chocolate. He then wanted to see (and taste) for himself the effect that tempering has on chocolate, particularly the effect that the cooling time had on the 'snap' of the chocolate when broken. Below is a snippet of the experimental method he used to test his idea.

Materials

- 1 × Kitchen scale
- 1 × Block of chocolate
- 1 × Medium sized pot
- 1 × Metal bowl
- 1 × Spatula
- 2 × Lined metal trays

Method

1. Chop chocolate into smaller blocks.
2. Label the trays A and B.
3. Add water into the pot until half full and bring the water to a simmer.
4. Once the water is simmering, turn off the stove.
5. Add chocolate into the metal bowl and place the bowl on top of the pot.
6. Leave for 1 minute until chocolate begins to melt.
7. Stir the chocolate with a spatula until melting.
8. Pour some of the melted chocolate onto the metal tray labelled A and use the spatula to spread the mix evenly.
9. Repeat for tray B.
10. Allow tray A to cool for 5 minutes, and tray B to cool for 10 minutes.
11. Break a piece of chocolate from each tray and record any observations relating to how easy it is to snap each sample.

a Identify the independent variable. (1 MARK)

b Identify the constant variables in the experiment. (2 MARKS)

c Warren noticed that when obtaining a sample of chocolate from tray A, the chocolate snapped very easily and cleanly whereas in tray B, the chocolate was still relatively soft. Explain how this result is possible. (3 MARKS)

Exam-style question hints

Q14a: The type of alloy depends on the size of alloying agents. **Q14b:** The force required to slide cations over each other and the electrostatic force can determine the hardness and strength of an alloy and a pure metal. **Q15a:** Choose the treatment that increases the softness and malleability of the metal. **Q15b:** Choose the treatment that increases the hardness and strength of the metal. **Q16a:** Consider factors such as scale and structure. **Q16b:** Materials used in energy-efficient buildings are those which can transfer a large amount of electricity. **Q17:** Both high surface area to volume ratio and appropriate temperature are important in drug delivery. **Q18a:** Iron makes up 90% of the total mass. **Q18b:** Carbon makes up 10% of the total mass. **Q19a:** Consider the SPDF configuration. **Q19b:** The trend of atomic size in the periodic table is determined by the difference in the positive nuclear charge of atoms. **Q20a:** An independent variable is the variable that the experimenter can control or change. **Q20b:** Constant variables are not changed between trials. **Q20c:** Consider the changes that happen to substances allowed to cool for a greater amount of time.

3C REACTIVITY OF METALS

In the previous lessons of this chapter, metals and their physical properties have been introduced. But how do they react with other substances or elements? How do their metallic characters determine their reactivity? This lesson will guide you through these concepts.

3A Metals | 3B Variations of metallic substances | 3C Reactivity of metals

Study design dot point

- experimental determination of the relative reactivity of metals with water, acids and oxygen

Key knowledge units

Reactions of metals with acids	1.1.6.1
Reactions of metals with water	1.1.6.2
Reactions of metals with oxygen	1.1.6.3

Key terms and definitions

Hydrogen pop test test used to indicate a reaction between a metal and an acid

Rust orange or brown metal oxide buildup on metals caused by oxygen

Metal oxide formed as a result of a reaction between a metallic element and oxygen

Lesson links

This lesson builds on:

- 2C The periodic table - part 1
- 2D The periodic table - part 2

Ionisation energy is related to the metallic bonding model, trends in atomic size can be used to explain the difference in properties of main group metals and transitions metals.

Reaction of metals with acids 1.1.6.1

OVERVIEW

Most metals can react with acids, however they will do so to different degrees.

THEORY DETAILS

Metal corrosion is one of the biggest problems caused by acid rain. If we recall from lesson 2C, metals tend to lose electrons easily due to their low ionisation energies. In a solution, an acid is separated into positive and negative ions. For example, hydrochloric acid HCl appears as H^+ and Cl^- ions in HCl solution. When a metal like magnesium (Mg) is placed into HCl solution, Mg atoms will give up their electrons and become positive Mg^{2+} ions. These Mg^{2+} ions will then be attracted by free negative ions in the solution Cl^- to form $MgCl_2$ which is called a salt. The remaining H^+ ions combine to form H_2 gas bubbles. We will learn more about ions in a later lesson. The reaction between Mg and HCl is demonstrated in figure 1.

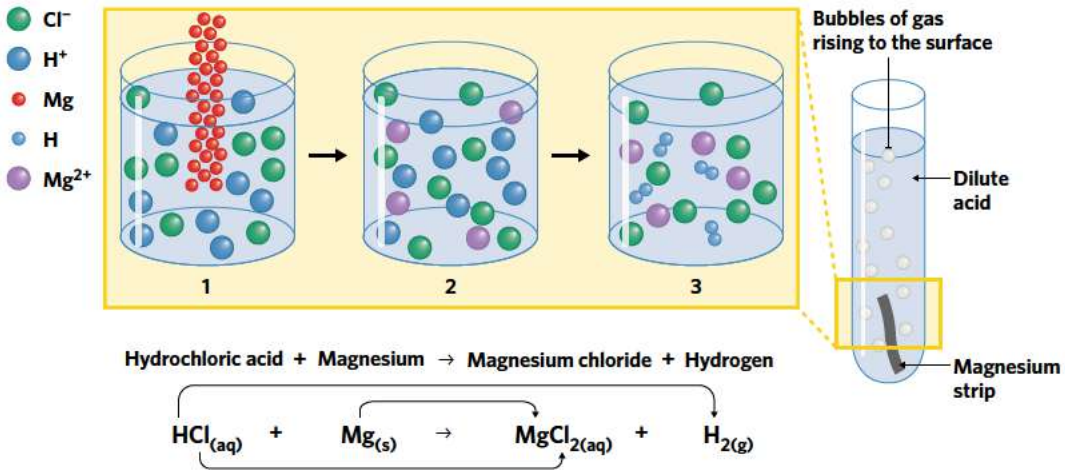


Figure 1 Demonstration of reaction between $Mg_{(s)}$ and $HCl_{(aq)}$.

By looking at the reaction equation between $Mg_{(s)}$ and $HCl_{(aq)}$ in figure 1, a general equation can be developed to show the reactions between metals and acids.

Acid + Metal → Ionic salt + Hydrogen gas

Since all reactions between metals and acids produce hydrogen gas, observed gas bubbles can be used to indicate that a reaction between a metal and acid has taken place. However, from lesson 2C, we know that some metals have higher ionisation energies than others so they react differently with acids. As an example, calcium (Ca) has higher ionisation energy than potassium (K). As a result, it is more difficult for Ca to lose electrons and the reaction of Ca with acids is weaker. The weaker the reaction is, the longer it takes to see the gas bubbles and therefore the reaction does not seem as vigorous. Therefore, how quickly reactions of metals and acid occur depends on the reactivity of the metals. The lower the ionisation energy of a metal is, the more easily it loses electrons, and the more rapidly it reacts with acid. Figure 2 demonstrates the trends in the reactivity of metals in the periodic table.

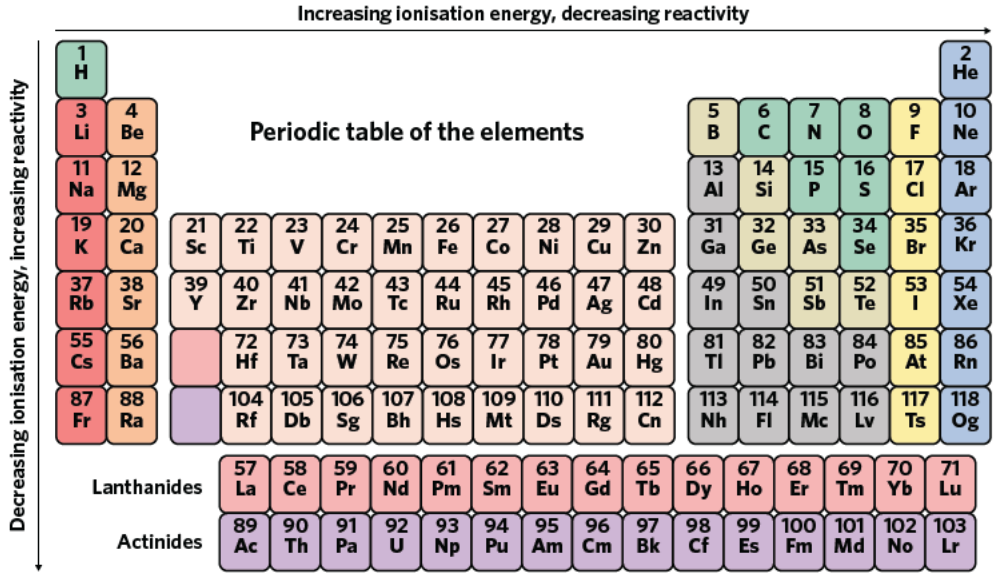


Figure 2 The trends of metal reactivity in the periodic table.

In the periodic table, as we go across a period from the left to the right, ionisation energy levels increase. This means that the reactivity of metals decreases and therefore the reactions with acids occur slower with less hydrogen gas bubbles produced. On the other hand, as we move down a group from the top to the bottom, ionisation energies levels decrease. Hence, the reactivity of metals increases and the reactions with acids occur faster with more hydrogen gas bubbles produced. How rapidly different metals react with acid is shown in figure 3.

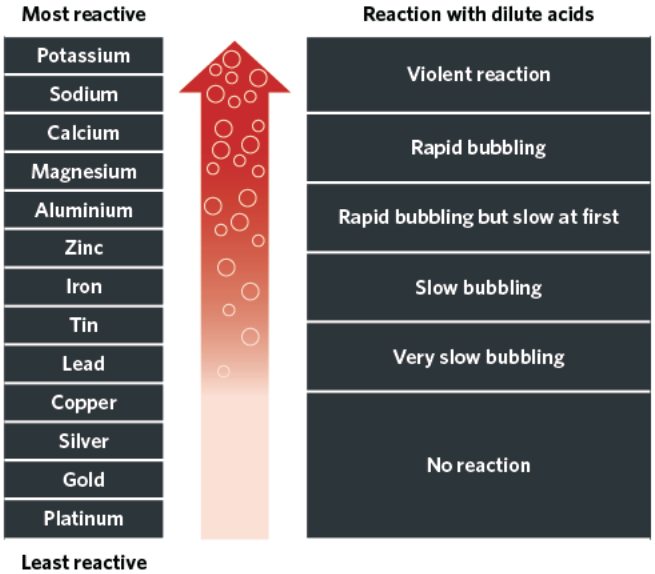


Figure 3 The trend in reactivity of different metals with acids.

Worked example 1

Does a reaction of tin with HCl or a reaction of sodium with HCl occur more quickly? Use the difference in ionisation energies of the two metals to justify your answer.

What information is presented in the question?

Two different metals.

What is the question asking us to do?

Determine the reaction of which metal with HCl occurs more quickly.

What strategy(ies) do we need in order to answer the question?

1. Compare the ionisation energies of the metals.
2. Use the difference in ionisation energies to compare the reactivity of the metals.
3. Use the difference in reactivity of the metals to determine which metal reacts more quickly with acid.

Answer

Sodium has lower ionisation energy than tin, so it is easier for sodium to lose electrons, meaning that sodium is more reactive than tin. Therefore, sodium reacts more quickly with HCl.

As reactions of some metals with acids occur slowly and do not produce enough hydrogen gas to see, a **hydrogen pop test** can be used to indicate metal-acid reactions. When a lit splint is placed inside a tube, a 'pop' sound will be heard if there is hydrogen gas produced from a metal-acid reaction. This is because hydrogen gas is highly flammable and therefore will readily react with oxygen in the air in the presence of a flame. The use of hydrogen pop test can be seen in figure 4.

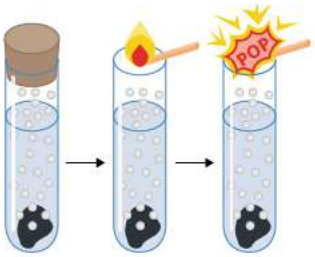


Figure 4 Hydrogen pop test.

Reaction of metals with water 1.1.6.2

OVERVIEW

Some metals can react with water while other metals cannot.

THEORY DETAILS

While metals play an important role in industry, there are some certain rules in storing and disposing of metals. This is due to the fact that metals in group 1 like lithium, sodium, potassium, rubidium and cesium (alkali metals) and metals in group 2 like calcium and magnesium can readily react with water as shown in figure 5.



Figure 5 Explosive reaction between sodium and water.

A general equation demonstrating a reaction of a metal with water can be generated as:

$\text{Metal} + H_2O \rightarrow \text{Metal hydroxide} + \text{Hydrogen gas}$

For example, a reaction between calcium and water will produce calcium hydroxide and hydrogen gas as shown:

$Ca_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$

However, not all metals react with water in the same way. Across a period in the periodic table from the left to the right, ionisation energy levels of metals increase, meaning that it is more difficult for the metals to lose electrons and therefore the reactivity of the metals with water decreases. For example, group 4 metals react slower with water compared to group 1 metals. In comparison, ionisation energy levels of metals decrease down a group, meaning that it is easier for them to lose electrons and therefore will react quickly with water. As an example, period 3 metals react faster with water than period 2 metals. The trend of reactivity of common metals with water can be seen in figure 6.

Metal	Reaction with water	Reactivity
Potassium	Violent with cold water	Most reactive
Sodium		
Calcium	Slow with cold water Rapid with steam	
Magnesium		
Aluminium	Usually no reaction	
Zinc		
Iron	Rusts slowly	
(Hydrogen)		
Copper	No reaction	
Silver		
Gold		Least reactive

Figure 6 The trend in reactivity of common metals with water.

Reaction of metals with oxygen 1.1.6.3

OVERVIEW

The orange and brown discolouration of metals is due to reactions between metals and oxygen.

THEORY DETAILS

Most of us would have come across **rust** which is the orange or brown discolouration built up on metals as shown in figure 7. But why and how does this happen? When exposed to oxygen over a certain period of time, metals react with oxygen to form inorganic compounds called **metal oxides**. For example, if a piece of iron is left in the atmosphere, a reaction between iron and oxygen will take place as shown in the chemical equation:

$2Fe_{(s)} + O_{2(g)} \rightarrow 2FeO_{(s)}$

The product of the reaction between iron and oxygen is called iron oxide as it consists of an iron ion and an oxide ion.



Image: Jackson Stock Photography/Shutterstock.com

Figure 7 Rust built up on a bicycle.

Most metals can react with oxygen except for silver (Ag), platinum (Pt) and gold (Au). The reactions between those that can react with oxygen have the general form as:

Metal + O_2 → Metal oxide

The trend of reactivity of metals with oxygen is similar to the trend of reactivity with water and acid. In the periodic table, the reactivity of metals with oxygen decreases as we go across a period from the left to the right, and increases as we move down a group. The trend of the reactivity of common metals with oxygen is shown in figure 8.

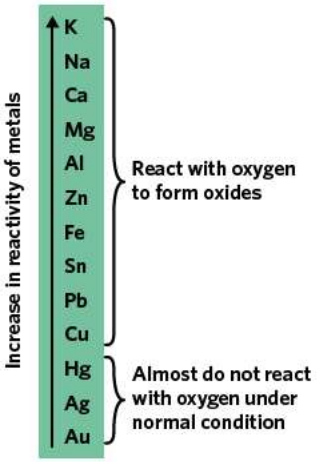
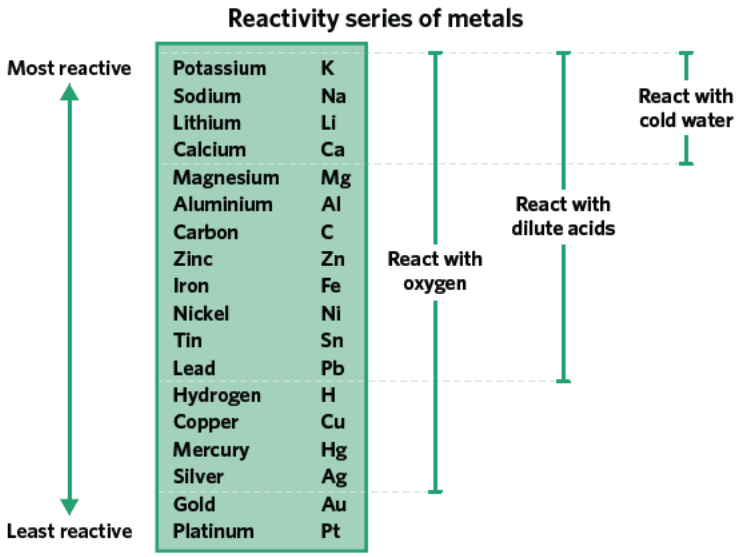


Figure 8 The trend of the reactivity of common metals with oxygen.

Theory summary

- A reaction between a metal and an acid will produce a salt and hydrogen gas.
- A hydrogen pop test can be used to detect reactions between metals and acids.
- A reaction between a metal and water will produce metal hydroxide and hydrogen gas.
- A reaction between a metal and oxygen will produce a metal oxide.
- Different metals react differently with acids, water and oxygen:



3C QUESTIONS

Theory review questions

Question 1

Most metals are able to react with acids.

A True

B False

Question 2

In a solution, metals tend to become positive ions because

A they tend to lose electrons.

B they tend to gain electrons.

Question 3

Across the groups from the left to the right in the periodic table, the reactivity of metals with acid

A increases.

B decreases.

Question 4

One of the products of a reaction between a metal and acid is

A oxygen gas.

B hydrogen gas.

Question 5

All metallic elements on the periodic table react with water at the same rate.

A True

B False

Question 6

A group 1 metal was placed into a beaker of water. The gas produced _____ produce a 'pop' sound in a hydrogen test.

A would

B wouldn't



Image: Sergey Merkulov/Shutterstock.com

Question 7

The product of a reaction between calcium and water is

A calcium hydroxide.

B calcium oxide.

Question 8

A toy truck has rust built up on it after being left exposed to the air for a year. This is the result of a reaction

A between a metal and carbon dioxide.

B between a metal and oxygen.



Image: Sergey Skleznev/Shutterstock.com

Question 9

All metals can react with oxygen under normal conditions.

A True

B False

Question 10

_____ can react with oxygen under normal conditions, whereas _____ cannot.

A Zinc, silver

B Silver, zinc

Deconstructed questions

Use the following information to answer questions 11–13.

Giao is a chemical engineer. While working in the laboratory on a humid day, she noticed that there was a small explosion with some gas produced on the shelf where she left some metals exposed to the air. No materials were burned by her colleagues during that time and the room temperature was maintained constant.

Question 11 (1 MARK)

What condition of the room can lead to the explosion?

A Humidity

B Temperature

C Atmospheric pressure

D All of above

Question 12 (1 MARK)

What types of metals may have been affected in the explosion?

A Group 3 metals with low ionisation energies

B Group 1 metals with low ionisation energies

C Group 2 metals with high ionisation energies

D Transition metals

Question 13 (3 MARKS)

Provide a possible cause of the small explosion and the presence of gas. Justify your answer including the type of metals that might have been affected in the explosion.

Exam-style questions

Within lesson

Question 14 (1 MARK)

Which of the following metals react with water at room temperature?

A Fe, Mg, Na, K

B Ni, Ag, Na, K

C Na, K, Ca, Li

D Hg, Cd, Na, K

Question 15 (3 MARKS)

Two boxes A and B contain two different metal pieces, calcium and lead. Thomson placed one piece of metal from each box into a tube containing dilute H_2SO_4 and observed the results shown in the table.

Box	Observation
A	Bubbles appear very slowly
B	More bubbles were produced at a faster rate

Identify the metals contained in box A and B. Justify your answer with reference to the difference in ionisation energies of the two metals.

Question 16 (5 MARKS)

A Chemistry student is given two jars of metals potassium and zinc which lost their name tags. The student suggests that the type of metals in each jar can be determined by placing 1 g of metal from each jar into a beaker of dilute sulfuric acid H_2SO_4.

a Is the student's method correct? Explain. (2 MARKS)

b Depending on the type of metals involved in a reaction, it is possible to distinguish between two metals based on their reactive nature. Based on what you learned in the lesson, suggest another method that can be used to determine the metal in each jar and include balanced chemical equation(s) in your answer. (3 MARKS)

Question 17 (2 MARKS)

An engineer wants to choose a metal to make agrimotors for farmers to use during the rainy season. Which metal, iron or aluminium, should the engineer choose? Explain your answer.

Multiple lessons

Question 18 (4 MARKS)

5.6 g of copper is placed into a solution containing an excess amount of nitric acid HNO_3.

$__Cu_{(s)} + __HNO_{3(aq)} \rightarrow __Cu(NO_3)_{2(aq)} + __H_{2(g)}$

a The equation given is unbalanced. Balance the equation provided by placing numbers in front of each species in the chemical reaction. (1 MARK)

b If the number of mol of copper used is equivalent to the amount of gas produced, calculate the mass of gas produced. (3 MARKS)

Question 19 (8 MARKS)

Ferrous ferric oxide Fe_3O_4 is the combination of iron (II) oxide and iron (III) oxide which can be produced by the following reactions:

$Fe_{(s)} + O_{2(g)} \rightarrow FeO_{(s)}$ (1)

$Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$ (2)

The formation of ferrous ferric oxide can be presented as:

$FeO + Fe_2O_3 \rightarrow Fe_3O_4$ (3)

a Balance equations (1) and (2). (2 MARKS)

b In an experiment, a student used 4.20 g for reaction 1 and 7.80 g for reaction 2.

i Calculate the mass of iron (II) oxide (FeO) produced if the amount produced is equivalent to the amount of iron used. (3 MARKS)

ii Calculate the mass of iron (III) oxide (Fe_2O_3) produced if the amount produced is equivalent to half the amount of iron used in reaction 2. (3 MARKS)

Key science skills

Question 20 (7 MARKS)

5 g of solid magnesium was mixed solid potassium and solid copper. A group of Chemistry students extracted pure solid magnesium from the mixture by the following steps:

1 Place cold water into the mixture and wait for 5 minutes.

2 Separate the remaining solid mixture and the solution produced.

3 Place concentrated sulfuric acid into the remaining solid mixture.

4 Collect the remaining solid which is pure magnesium.

a Identify the metals in the solid mixture in step 2. Justify your answer including the occurred chemical reaction(s). (3 MARKS)

b One student in the group suggested that sulfuric acid should be placed into the mixture first instead of the cold water. Is the student's suggestion correct? Explain your answer. (2 MARKS)

c After finishing all four steps, the students weighed the collected sample of magnesium and realised that they could only collect 4.3 g magnesium. Explain why this is the case and suggest one change that the students can make to increase the amount of pure magnesium extracted. (2 MARKS)

Exam-style question hints

14: Only group 1 and 2 metals react with water at room temperature. **15:** Different metals have different ionisation energies. **16a:** Two different metals can be determined if different observations are obtained from their reactions with the same substance. **16b:** Choose a substance that the metals react with differently. **17:** Rain water reacts differently with the metals. **18a** The number of each element needs to be balanced on both the reactant and product side of a chemical reaction. **18b:** The mass of the gas produced depends on both the molar mass and amount of substance present. **19a** The number of each element needs to be balanced on both the reactant and product side of a chemical reaction **19b:** The mass of the gas produced depends on both the molar mass and amount of substance present. **20a:** The reactivity of the metals with water is different. **20b:** The purpose of the experiment is to extract pure magnesium. **20c:** Determine in which step magnesium has been lost.

EXPERIMENT

EXPLORING THE REACTIVITY OF GROUP 2 METALS

In this chapter, we have explored the different properties of metals. This experiment will investigate and compare the reactivity by mass of two different metals belonging to Group 2 of the periodic table, magnesium and calcium.

Materials

- 2 × Test tubes
- 50 mL 0.05 M HCl
- 1.0 g strip of magnesium ribbon
- 1.0 g piece of calcium
- 1 × Test tube rack
- 2 × 2.0 mL pasteur pipette
- 1 × Timer



Coil of magnesium ribbon



Pieces of calcium

Image (left to right): Stephen Barnes & Bjoern Wylezich/Shutterstock.com

Method

1. Label each test tube A and B.
2. Fill each test tube roughly halfway with the 0.05 M HCl using the pasteur pipette.
3. Place the strip of magnesium ribbon into test tube A.
4. Start the timer and record the time it takes for bubbles to stop appearing.
5. Record this time and any additional observations.
6. Repeat steps 3–5 for test tube B using the sample of calcium.

Results

Metal	Time taken for bubbles to stop appearing (seconds)	Observations
Magnesium		
Calcium		

QUESTIONS

Question 1 (1 MARK)

Identify the independent and dependent variables in this experiment.

Question 2 (2 MARKS)

In order for the experiment to compare the reactivity of magnesium and calcium in a valid manner, what variable(s) need to be controlled?

Question 3 (2 MARKS)

Are there any flaws in the experimental design? Suggest a change to the experimental design that would resolve any of these flaws.

Question 4 (6 MARKS)

a Compare the reactivity of magnesium and calcium based on the results collected. (1 MARK)

b A student generated the following table of results:

Metal	Time taken for bubbles to stop appearing (seconds)	Observations
Magnesium	32	Steady stream of bubbles, metal slowly dissolves
Calcium	16	Rapid stream of bubbles, test tube becomes hot

i Are the results expected? Justify your answer with reference to trends in the periodic table. Explain a possible cause for any discrepancies. (3 MARKS)

ii Comment on the reproducibility of the results collected. (2 MARKS)

Question 5 (2 MARKS)

Identify two safety considerations when handling hydrochloric acid and reactive metals.

ANSWERS

1 Independent variable: Identity of metal (magnesium or calcium)
Dependent variable: Time taken for a complete reaction to occur

2 The following variables should be controlled:

- Surface area of metal
- Concentration of HCl
- Purity of metal
- Volume of HCl
- Timer

3 [According to the current experimental design, the volume of hydrochloric acid required in each test tube is not specified.[1]][This affects the validity of the experiment because the time taken to react completely also depends on the amount of hydrochloric acid present and, therefore, there are multiple independent variables.[2]][The validity of the experiment could be improved by specifying the volume of acid to be used for consistency.[3]]

✓ ✗ I have identified that the amount of acid required is not specified.[1]

✓ ✗ I have explained its effect on validity.[2]

✓ ✗ I have identified an improvement to resolve the validity of the experiment.[3]

4 **a** *Please check your answer with your teacher.

b **i** [For metals, reactivity is a measure of how easily they are able to lose their valence electron(s).[1]][Although both elements belong in the same group, calcium is further down the group than magnesium. Calcium's valence electrons are further away from the nucleus, and therefore, are attracted less strongly to the nucleus. This means they can be lost more easily and, therefore, calcium is more reactive than magnesium.[2]][Therefore, it is expected that there will be a more rapid stream of bubbles from the calcium test tube due to calcium reacting faster than magnesium.[3]]

✓ ✗ I have defined reactivity.[1]

✓ ✗ I have compared the reactivity of magnesium and calcium.[2]

✓ ✗ I have compared the reactivity of magnesium and calcium based on the observations.[3]

ii [Reproducibility is a measure of whether a different experimenter can generate similar results based on the experimental design.[1]][This experiment is not very reproducible because the volume of hydrochloric acid required is not specified, and subjective observations like temperature and sight will vary between individuals.[2]]

✓ ✗ I have defined reproducibility.[1]

✓ ✗ I have identified the reproducibility of the experiment and described the factors that may affect the reproducibility of the experiment.[2]

5 [Hydrochloric acid should be handled with protective eye-wear, a lab coat and gloves.[1]][Reactive metals like calcium and magnesium should only be used in small quantities and kept away from ignition sources, water and concentrated acids.[2]]

✓ ✗ I have identified safety considerations when handling acids.[1]

✓ ✗ I have identified safety considerations when handling reactive metals.[2]

CHAPTER 3 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Question 1 (1 MARK)

In which group on the periodic table would we find the most reactive metals?

A Group 17
B Group 4
C Group 2
D Group 1

Question 2 (1 MARK)

Sodium, rubidium and caesium are all alkali metals. Which of the following options orders the metals in increasing order of reactivity?

A Rubidium, sodium, caesium
B Sodium, caesium, rubidium
C Caesium, rubidium, sodium
D Sodium, rubidium, caesium

Question 3 (1 MARK)

In chemistry, rust is also referred to as a

A metallic hydrolyte.
B metal oxide.
C dihydrogen oxide.
D metalloid.

Question 4 (1 MARK)

Gold jewellery is produced by heating solid gold to a temperature of 1050°C and allowing it to melt so that it can be poured into different shaped moulds. This suggests that metallic substances

A are highly combustible.
B are expensive.
C contain strong bonds.
D are easily malleable.

Question 5 (1 MARK)

A company wanted to develop an interstitial alloy to be used for their newest car model. They wanted to use lithium as part of the alloy, which has an atomic radius of approximately 146 picometers (pm). Which of the following elements would be most suitable for the production of a substitutional alloy?

	Element A	Element B	Element C	Element D
Atomic radius (pm)	143	175	200	100

A Element A
B Element B
C Element C
D Element D

Question 6 (1 MARK)

Which of the following materials should not be composed of metallic elements?

A Gloves used by electricians to fix live wires.
B Wires used in telephone communications.
C Cooking pots.
D Light bulb filaments.

Question 7 (1 MARK)

In a reaction between a strong acid and potassium metal, it is expected that we would

A smell a fruity odour.
B observe bubbles.
C observe a change in the colour of the acid.
D see no change.

Question 8 (1 MARK)

A metal was treated under a certain condition to form this relatively large crystal pattern. The metal was most likely

A quenched.
B annealed.
C hammered.
D tempered.



Image: Matteo Galimberti/Shutterstock.com

Question 9 (1 MARK)

Which of the following metals is least likely to react with water at room temperature?

A Magnesium
B Rubidium
C Copper
D Beryllium

Question 10 (1 MARK)

Properties such as lustre and heat conductivity of metals are mainly due to

A the reactivity of metals.
B the size of metallic atoms.
C delocalised electrons.
D electronegativity of metals.

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (11 MARKS)

A large mining company, Sparks Industries, based in Perth wants to develop a new iron ore mine and extraction facility in the Pilbara region of Western Australia. In order to get the proposal approved, there is a complex consultation process which involves many key stakeholders, including scientists, environmentalists and locals, that takes place in the town hall.

a What is an ore and what happens to ore before a metal is extracted? (2 MARKS)

b The scientist begins to speak about how iron is extracted from its ore. Describe the processes that take place and the materials used. Equations are not required. (3 MARKS)

c The environmentalist describes a large-scale impact of iron production and a local impact. Identify the two possible impacts that the environmentalist could be referring to. (2 MARKS)

d The locals are split into two groups, those who support the mine and those who are against. With reference to the societal impacts of mining, why could this be the case? (2 MARKS)

e Iron has a variety of different uses. With reference to the properties of metals, explain why iron is often used in weight bearing structures like bridges. (2 MARKS)

Question 12 (10 MARKS)

A materials scientist wants to create a new type of alloy to use as a wire in a heart pacemaker. Heart pacemakers send electrical pulses to the heart through wires inserted into the body's veins as shown in the image. Therefore, the scientist needs the alloy to possess the following properties:

- Able to conduct electricity
- Malleable
- Ductile
- Durable

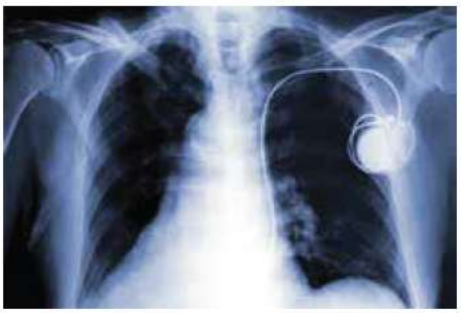


Image: khuruzero/Shutterstock.com

a What is an alloy? (1 MARK)

b Identify the two main types of alloys and their defining characteristics. (2 MARKS)

c With reference to metallic bonding, explain why metals are able to conduct electricity? Use a diagram to explain your answer. (3 MARKS)

d Why would the wires in a heart pacemaker need to be malleable and able to conduct electricity? (1 MARK)

e The materials scientist wants to further modify her alloy and is considering either annealing, quenching or tempering. Which technique would be most suitable? Justify your answer with reference to recrystallisation. (3 MARKS)

Question 13 (9 MARKS)

Frank has been given three pieces of grey metal which look identical. He knows that one metal is magnesium, one metal is barium and one metal is beryllium. He wants to determine the identity of each metal based on their chemical reactions and properties.

a Which metal would be expected to have the greatest reactivity? Justify your answer with reference to trends in the periodic table. (2 MARKS)

b Write the equation of the chemical reaction that would occur between solid barium and water. If this reaction initially occurred in a sealed test tube, what would happen when a flame is placed inside the test tube after removing the seal? (2 MARKS)

c Frank proposes that he could also determine the reactivity series by leaving the metals out in the open air. What would happen to the metal samples over time (include any relevant equations) and how could this be used to differentiate between them? (5 MARKS)

KEY SCIENCE SKILLS QUESTIONS (6 MARKS)

Question 14 (6 MARKS)

Caesium is a highly reactive metal and is a liquid at room temperature. One of its applications is in atomic clocks which are the primary standards for global timekeeping. A scientist was asked to repair an atomic clock which was being kept in a vacuum with no oxygen present.

a Why would liquid caesium be kept in a vacuum? Provide any relevant chemical equations. (2 MARKS)

b Before handling caesium, the scientist practices her technique using sodium. Is this a valid way to practice? (2 MARKS)

c Atomic clocks are very accurate. Do they have to be precise in order to be an effective timekeeping method? (2 MARKS)

UNIT 1 AOS 1, CHAPTER 4

Ionic Compounds

04

4A Ionic compounds

4B Properties of ionic compounds

4C Writing formulas for ionic compounds

Key knowledge

- common properties of ionic compounds (brittleness, hardness, high melting point, difference in electrical conductivity in solid and liquid states) with reference to their formation, nature of ionic bonding and crystal structure including limitations of representations
- experimental determination of the factors affecting crystal formation of ionic compounds
- the uses of common ionic compounds
- precipitation reactions represented by **balanced, full** and ionic **equations**, including states

Image: Lubo Ivanko/Shutterstock.com

4A IONIC COMPOUNDS

In the previous chapter, we learned how salts and oxides could be produced from reactions of metals with acids and oxygen, respectively. In this lesson, we will further explore these compounds including their formation and their model.

4A Ionic compounds | 4B Properties of ionic compounds | 4C Writing formulas for ionic compounds

Study design dot points

- experimental determination of the factors affecting crystal formation of ionic compounds
- common properties of ionic compounds (brittleness, hardness, high melting point, difference in electrical conductivity in solid and liquid states) with reference to their formation, nature of ionic bonding and crystal structure including limitations of representations

Key knowledge units

Formation of ionic compounds	1.1.11.1
The ionic bonding model	1.1.10.1

Key terms and definitions of this lesson

Ion charged atom

Cation positive ion formed when an atom loses it valence electron(s)

Anion negative ion formed due to the addition of electron(s) to an atom

Ionic bond electrostatic force of attraction between cations and anions

Ionic compound compound made up of ions held together by electrostatic forces of attraction

Lesson links

This lesson builds on:

▶ 2C The periodic table - part 1

The formation of cations and anions is due to the ionisation energy levels and electronegativity of metals and non-metals.

▶ 3C Reactivity of metals

The formation of ions and ionic compounds can be seen in reactions of metals with acids and oxygen.

Formation of ionic compounds 1.1.11.1

OVERVIEW

The majority of ionic compounds are formed from metal atoms and non-metal atoms.

THEORY DETAIL

Recall from chapter 2, we know that metal atoms have low ionisation energy, meaning that it does not require much energy to remove valence electron(s) from metal atoms to obtain a full valence shell. Therefore, metal atoms tend to lose electrons to form positive **ions** called **cations.** For example, a sodium atom (Na) tends to lose one electron in its valence shell to form a sodium ion with a positive charge as shown in figure 1.

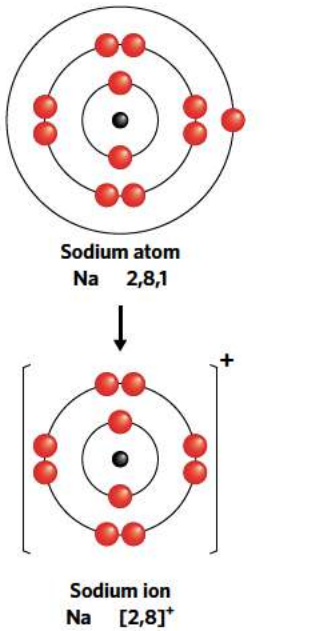


Figure 1 Formation of a sodium ion from a sodium atom.

Non-metal atoms tend to attract electrons due to their degree of electronegativity. An electron(s) is added to the outer shell of the non-metal atoms to form a full valence shell. This results in the formation of negatively charged ions called **anions.** For example, a chlorine atom (Cl) has a tendency to attract one electron to its valence shell to form a chloride ion with a negative charge as shown in figure 2.

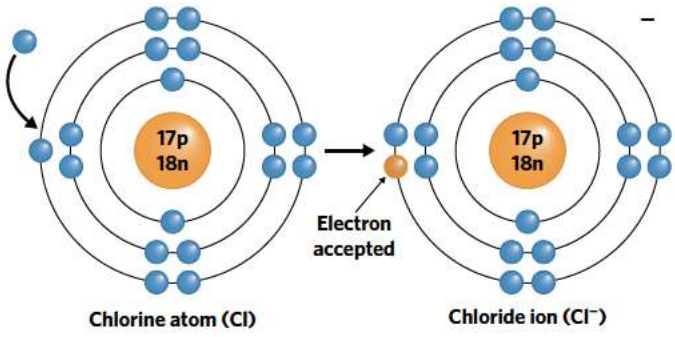


Figure 2 Formation of a chloride ion from a chlorine atom.

When writing the symbol for an ion, the charge is written as a superscript after the chemical symbol as shown in figure 3.

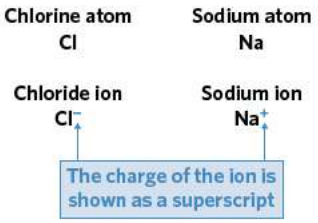
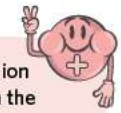


Figure 3 Representation of ions.

Due to the fact that metal atoms tend to lose electrons while non-metal atoms tend to attract electrons, when a metal reacts with a non-metal, the electron(s) removed from the metal can be accepted by the non-metal. This results in the formation of two oppositely charged ions; a cation and anion. The oppositely charged ions become attracted to each other and form an **ionic bond**. An ionic bond is the electrostatic force of attraction between a cation and an anion. It's important to note that ionic bonding is formed only when the difference in electronegativity between two elements is greater than or equal to approximately 2. This results in the formation of an **ionic compound**, commonly referred to as a salt.



If the charge of the ion is more than 1, then the numerical value of the charge is written before the symbol and is also superscripted. Eg. Mg^{2+}

An ionic compound can be formed from a sodium ion and a chloride ion as demonstrated in figure 4. One electron removed from the valence shell of the sodium atom is accepted by the chlorine atom into its valence shell. Since the sodium ion (Na^+) becomes positively charged and the chloride ion (Cl^-) becomes negatively charged, there is an electrostatic force that holds them together. As a result, an ionic compound sodium chloride (NaCl) is able to form.

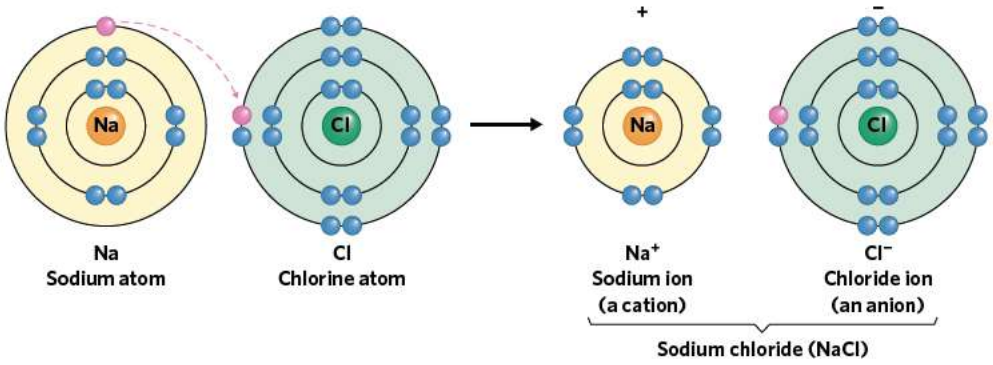


Figure 4 Formation of sodium chloride from a sodium atom and a chlorine atom.

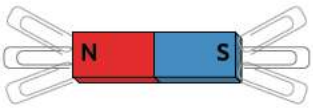


Figure 5 Demonstration of a magnet's ability to attract multiple objects.

Since ionic compounds are formed based on the forces of attraction between opposite charges, similar to the way a pole of a magnet can attract multiple objects as shown in figure 5, there can be many charged particles attracted to the positive or negative ion as illustrated in figure 6.

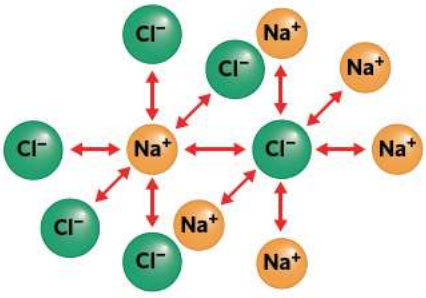


Figure 6 Illustration of the ability of a positive or negative ion to attract multiple charged particles.

The ratio of ions in an ionic compound is affected by the number of electrons that can be lost by a metal and accepted by a non-metal. As shown in figure 7, the magnesium atom (Mg) loses two electrons from its valence shell to obtain the octet rule while a chlorine atom (Cl) can only accept one electron to obtain the octet rule. Therefore, the electrons lost by the magnesium atom can be received by two chlorine atoms.

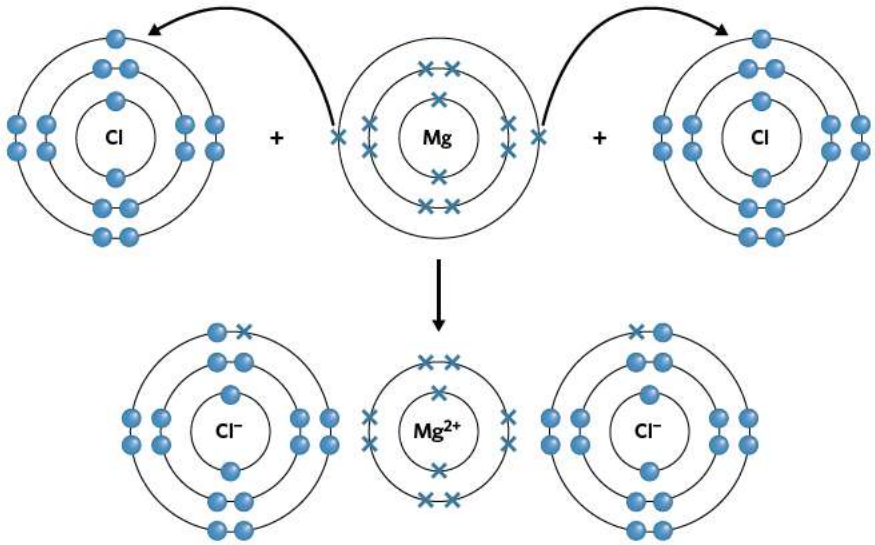


Figure 7 Formation of magnesium and chloride ions from a magnesium atom and two chlorine atoms.

Worked example 1

Use Bohr models to explain the formation of the ionic compound Al_2S_3.

What information is presented in the question?

Ionic compound Al_2S_3.

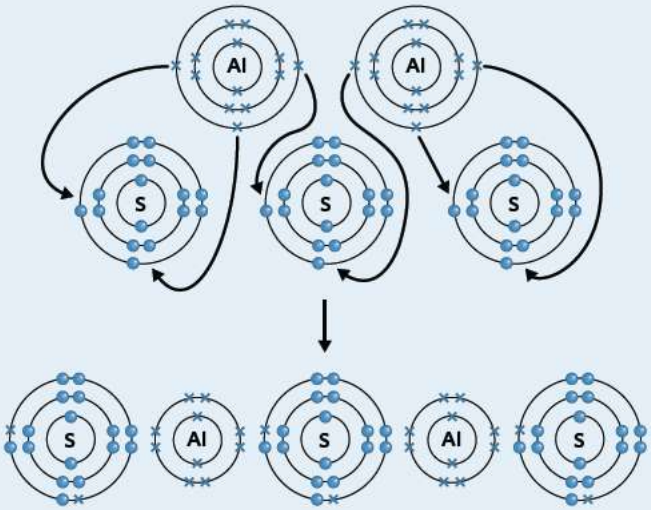
What is the question asking us to do?

Use Bohr models to explain the formation of the ionic compound.

What strategy(ies) do we need in order to answer the question?

1. Write the Bohr model of one Al atom.
2. Write the Bohr model of one S atom.
3. Determine the number of electrons that can be removed from one Al atom.
4. Determine the number of electrons that can be accepted by one S atom.
5. Determine the number of Al and S atoms required.
6. Show the formation of Al_2S_3 from Al atoms and S atoms.

Answer



The ionic bonding model 1.1.10.1

OVERVIEW

Ionic bonding in ionic compounds can be represented in different ways.

THEORY DETAILS

Bohr models

The Bohr models allow us to see how ionic bonds are formed and the ratio of cations and anions, however it does not show how ions are arranged in a 3D space.

3D crystal lattices

An ionic compound is actually a three-dimensional structure consisting of multiple ions bonding to each other rather than a flat two-dimensional structure of a single cation and a single anion. As discussed previously, this is due to the fact that ionic bonds can be formed between a single cation and multiple anions if they are close enough to each other. Similarly, ionic bonds can be formed between a single anion and multiple cations. Many cations and anions in an ionic compound are bonded to each other forming a three-dimensional structure called a **crystal lattice.**

As presented in figure 8, the crystal lattice of sodium chloride contains multiple sodium ions and chloride ions bonded to each other by the electrostatic force. The more ions that are bonded to each other, the stronger the electrostatic force and therefore the more stable the lattice. Therefore, to maximise the stability of the lattice, a sodium ion is bonded to six surrounding chloride ions, as shown in figure 8. It's important to note that the bonds between each sodium atom and its six neighbouring chlorine atoms are all ionic bonds. Similarly, a chlorine ion is bonded to six neighbouring sodium ions via ionic bonds.

Because chloride ions are close to each other and sodium ions are close to each other, there is repulsive force between chloride ions and between sodium ions. However, the electrostatic force between sodium ions and chloride ions is much stronger than the repulsive force, meaning that the structure of the crystal lattice is relatively stable.

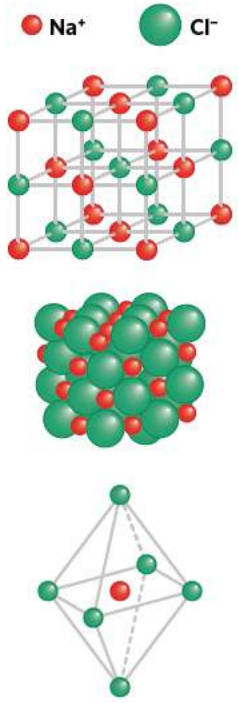


Figure 8 Demonstration of the crystal lattice of sodium chloride.

Although crystal lattices of ionic compounds allow us to visualise how ions are arranged in three dimensional space, it still has some limitations:

- The distance between ions and the sizes of ions are not to scale.
- It does not provide information about how electrons are removed from atoms and how electrons are accepted by atoms.
- It does show how ionic bonds are formed.

Factors affecting the size of ionic crystals

Crystals of an ionic compound can be collected from its solution by performing the following steps:

- Heat a solution containing the dissolved ionic compound.
- Allow the heated solution to cool down and allow the liquid to evaporate.
- Collect the crystals of the ionic compound.

As the heated solution cools down and the liquid evaporates, the dissolved ions are bonded to each other to form crystals. The size of ionic crystals depends on how long the solution is left to cool down and how long the liquid is left to evaporate. The longer the solution is left to cool down and the liquid is left to evaporate, the larger the crystals that can be formed because they have more time to grow. For example, larger crystals will be formed in a heated solution left to cool down for five minutes in comparison to a heated solution left to cool down for four minutes. This is similar to how larger metallic crystals are formed when melted metals are left to cool down more slowly in lesson 3B.

Theory summary

- Most ionic compounds are formed from metal atoms and non-metal atoms.
- The ratio of atoms of each element in an ionic compound depends on the number of electrons removed by a metal and accepted by a non-metal.
- In an ionic compound, the bond between metal ion(s) and non-metal ion(s) is called ionic bonding.
- A single ion is surrounded by (and therefore bonded to) multiple oppositely charged ions.
- Ionic compounds are arranged as a crystal lattice.
- The strength of the ionic bond is greater than the forces of repulsion of ions of the same charge.
- Ionic compounds that are allowed to cool for longer periods of time will result in larger crystals.

4A QUESTIONS

Theory review questions

Question 1

Barium can lose two valence electrons to form a cation. Barium ion can be represented as

A $2Ba^{+}$.

B Ba^{2+}.

Question 2

An ionic bond is formed when there's electrostatic force between

A a metal atom and a non-metal atom.

B a positive metal ion and a negative non-metal ion.

Question 3

An ionic bond can be formed from any metals and non-metals.

A True

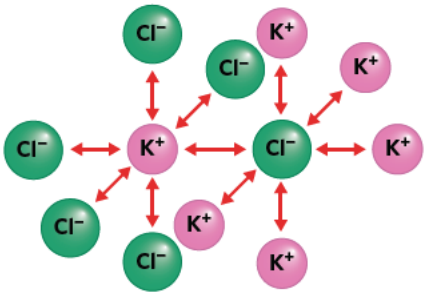
B False

Question 4

Which representation best illustrates an ionic compound?

A Representation 1

B Representation 2



Representation 1



Representation 2

Question 5

Bohr models can show _____ but cannot demonstrate _____.

A how ionic bonds are formed, how ions are arranged in space

B the ratio of cations and anions, how ionic bonds are formed

Question 6

In a sample of an ionic compound, each cation is only bonded to the anion(s) it donates its electrons to.

A True

B False

Question 7

A crystal lattice of an ionic compound is _____ structure because _____.

A a flat 2D, each cation is only bonded to the non-metal ion(s) it donates its electrons to

B a 3D, a single cation can be bonded to multiple anions if they are close enough to each other

Question 8

In a sample of sodium chloride, each sodium ion is bonded to _____ chloride ions and each chloride ion is bonded to _____ sodium ions.

A 6, 6

B 6, 4

Question 9

In a sample of sodium chloride, the bond between a sodium ion and the chloride ion it donates its electron to is an ionic bond, while the bond between that sodium ion and the other five chloride ions is another type of attractive force.

A True

B False

Question 10

In making crystals of an ionic compound, crystals can grow bigger as they are left to cool down more _____ because _____.

A slowly, they have more time to grow

B quickly, the crystals will be dissolved again over time

Deconstructed questions

Use the following information to answer questions 11–13

Santana, a Chemistry student, was demonstrating the ionic bonds present in calcium fluoride (CaF_2). The model shown was her demonstration. She explained that one Ca^{2+} is only bonded to two F^- ions and one F^- ion is only bonded to one Ca^{2+} ion.



Question 11 (1 MARK)

An ionic bond is _____ between a positive _____ ion and a negative _____ ion. A single positive ion can be bonded to _____. A single negative ion can be bonded to _____.

A an electrostatic force, non-metal, metal, one negative ion, multiple positive ions

B mechanical force, metal, non-metal, multiple negative ions, one positive ion

C an electrostatic force, metal, non-metal, multiple negative ions, multiple positive ions

D mechanical force, metal, non-metal, multiple negative ions, multiple positive ions

Question 12 (1 MARK)

Was the student's demonstration correct?

A Correct, because one F^- ion is only bonded to one Ca^{2+} ion.

B Incorrect, because one Ca^{2+} is only bonded to one F^- ion, not two.

C Correct, because all ions have obtained the octet rule.

D Incorrect, because one F^- ion is bonded to multiple Ca^{2+} ions.

Question 13 (3 MARKS)

Evaluate the student's demonstration. Suggest a model that best represents ionic bonds present in calcium fluoride (CaF_2).

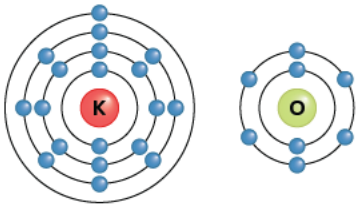
Exam-style questions

Within lesson

Question 14 (3 MARKS)

The Bohr models of potassium and oxygen are shown.

Use the given Bohr models to demonstrate the formation of K_2O.



Question 15 (6 MARKS)

In veterinary medicine, KBr is widely used to treat epilepsy in animals.

a Draw Bohr models to demonstrate the formation of a potassium ion from a potassium atom. (2 MARKS)

b Draw Bohr models to demonstrate the formation of a bromine ion from a bromine atom. (2 MARKS)

c Draw Bohr models to show how KBr is formed from the ions in questions a and b. (2 MARKS)

Question 16 (4 MARKS)

The electronegativity of different elements are shown in the table.

Element	Electronegativity
Ca	1.00
C	2.55
Cl	3.16

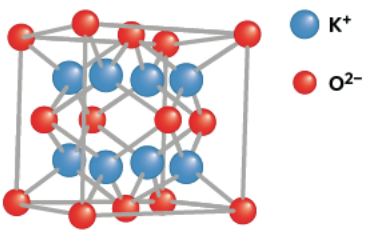
a Can an ionic bond be formed between calcium and chlorine? Explain your answer. (2 MARKS)

b Can an ionic bond be formed between carbon and chlorine? Explain your answer. (2 MARKS)

Question 17 (3 MARKS)

The 3D structure shown is the lattice of potassium oxide (K_2O) in which each potassium ion is bonded to four neighbouring oxygen ions.

A student stated that the bond between each potassium ion and the oxide ion it donates its electrons to is an ionic bond, while the bonds between that potassium ion and the other three neighbouring oxide ions are not ionic bonds. Is the student's statement accurate? Justify your answer.



Multiple lessons

Question 18 (4 MARKS)

Brooke wants to make magnesium chloride from magnesium and hydrochloric acid using the equation:

$Mg + HCl \rightarrow MgCl_2 + H_2$

a Balance the given chemical equation. (1 MARK)

b What type of compound is magnesium chloride? Explain your answer. (2 MARKS)

c If Brooke uses 7.8 g of magnesium, what is the number of moles of magnesium used in the reaction? (1 MARK)

Question 19 (7 MARKS)

Both sodium (Na) and sodium chloride (NaCl) have lattice structures. However, they have different melting points as shown in the table.

Lattice	Melting point (°C)
Na	98
NaCl	801

a What chemical bond is present in each lattice? (2 MARKS)

b Based on the difference in melting points shown in the table, which of the chemical bonds determined in question 19a is stronger? Justify your answer. (2 MARKS)

Key science skills

Question 20 (5 MARKS)

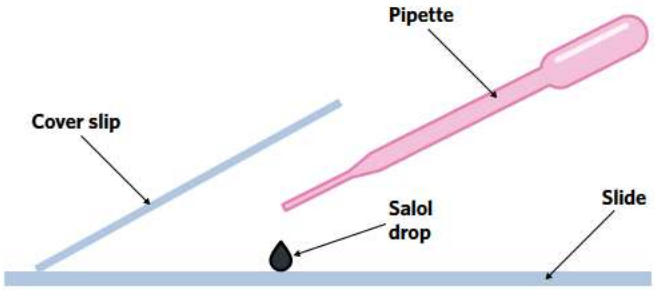
Mikkel conducted an experiment to investigate the formation of ionic crystals.

Materials

- 1 × Microscope
- 2 × Dry microscope slides (1 heated and 1 cold)
- 2 × Cover slips
- 1 × Pipette
- Paper towel
- 1 × Stopwatch
- Safety glasses
- 4 g melted salol (phenyl salicylate)

Method

1. Put a drop of melted salol on the cold slide.
2. Cover the slide with a cover slip.
3. Place the slide under the microscope.
4. Set stopwatch to 0.
5. Start the stopwatch when starting looking at the drop under the microscope.
6. Note the time when crystals start to be formed.
7. Note the time when no more crystals are being formed.
8. Repeat the previous steps for the heated slide.
9. Place each slide under the microscope again to compare the size of crystals.



Mikkel recorded his observations in the shown table.

Slide	Time when crystals start to be formed (seconds)	Time no more crystals are formed (seconds)
Cold	5	25
Heated	20	100

a Identify independent and dependent variables in the experiment. (2 MARKS)

b On which slide would Mikkel see larger crystals? Explain your answer. (3 MARKS)

Exam-style question hints

14: Metals donate electrons to non-metals in such a way that both metals and non-metals obtain the octet rule. **15:** Difference in electronegativity between a metal and a non-metal must be larger than approximately 2 for an ionic bond to be formed. **16:** The number of ionic bonds formed depends on the proximity of ions to other ions. **17:** Metals donate electrons to non-metals in such a way that both metals and non-metals obtain the octet rule. **18a:** The number of elements are the same at both sides of the equation. **18b:** The type of compound can be determined based on the elements present. **19a:** Stronger bonds require more heat to be broken. **19b:** Electrical conductivity requires the movement of electrons. **20a:** Dependent variables are the variables being measured or tested. **20b:** The size of crystals formed depends on how long they are allowed to be formed.

4B PROPERTIES OF IONIC COMPOUNDS

In the previous lesson, we learned about the formation of ionic compounds and the nature of ionic bonds. In this lesson, we will explore how ionic bonds determine properties of ionic compounds and the applications of ionic compounds.

4A Ionic compounds | 4B Properties of ionic compounds | 4C Writing formulas for ionic compounds

Study design dot points

- common properties of ionic compounds (brittleness, hardness, high melting point, difference in electrical conductivity in solid and liquid states) with reference to their formation, nature of ionic bonding and crystal structure including limitations of representations
- the uses of common ionic compounds

Key knowledge units

Properties of ionic compounds	1.1.10.2
Uses of ionic compounds	1.1.12.1

Key terms and definitions

Brittle tendency of a material to fracture (break) when subjected to pressure

Hardness resistance to deformation when subjected to pressure

Melting point temperature at which a substance changes its state from solid to liquid

Soluble able to be dissolved in a given solvent

Insoluble unable to be dissolved in a given solvent

Lesson links

This lesson builds on:

- 4A Ionic compounds

Properties of ionic compounds are determined by the nature of ionic bonding and ionic crystal lattice.

Properties of ionic compounds 1.1.10.2

OVERVIEW

Properties of ionic compounds are determined by the characteristics of ionic bonds and 3D ionic crystal lattices.

THEORY DETAIL

Brittleness and hardness

As we learned in lesson 4A, the ionic bonds between positive and negative ions within crystal lattices are strong, meaning that it requires a large amount of force to disrupt these bonds. Therefore, ionic compounds have a high degree of **hardness**.
However, when a strong enough force is applied to an ionic compound, ions with the same charge will be pushed closer to each other. The repulsive force between like-charged ions is strong enough to shatter and break the lattice. As a result, ionic compounds are described as being **brittle**. As shown in figure 1, when the crystal lattice of sodium chloride is struck with a hammer, the lattice is disrupted in a way that allows negative chloride ions to be closer to each other. This creates a strong repulsive force between negative chloride ions, which is able to split the lattice.

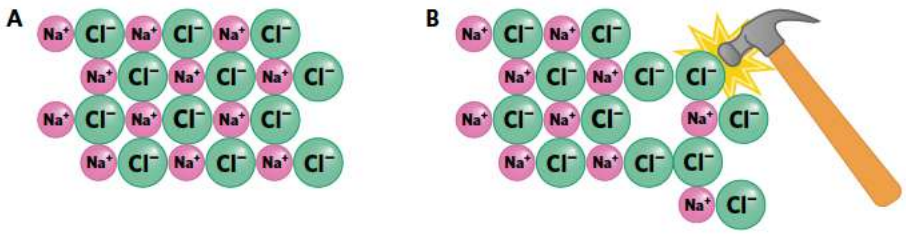


Figure 1 Demonstration of the brittleness of sodium chloride when subjected to a strong mechanical force.

Melting point

Due to the fact that the ionic bonds between positive and negative ions within crystal lattices are strong, it requires a lot of energy to break these bonds. Hence, ionic compounds have relatively high **melting points** compared to other substances. For example, the melting point of MgO is as high as 2852°C.

Electrical conductivity

The ability of ionic compounds to conduct electricity depends on their physical states. For a substance/material to conduct electricity, it needs to allow for charged particles to move through to carry an electric current. As we learned in lesson 4A, cations and anions are in fixed positions in an ionic lattice, meaning there is no movement of charged particles. Therefore, an ionic compound in solid state is not able to conduct electricity as shown in figure 2A.

However, if an ionic compound is molten (liquid state) or dissolved in water (aqueous state), its ionic lattice is broken, which results in individual cations and anions that are able to move freely. Hence, when connected to an electrical circuit, the ionic compound in liquid or aqueous state can conduct electricity. As shown in figure 2B, when an ionic compound is dissolved in water, individual cations and anions can move around in the solution. Consequently, the ions are able to carry electric charge through the solution, allowing the electric current to continue moving around the circuit. Also illustrated in figure 2B, this allows the light bulb in the circuit to shine, demonstrating that the ionic compound can conduct electricity when in aqueous form. In general, the higher the concentration of the ionic compound, the higher the electrical conductivity.

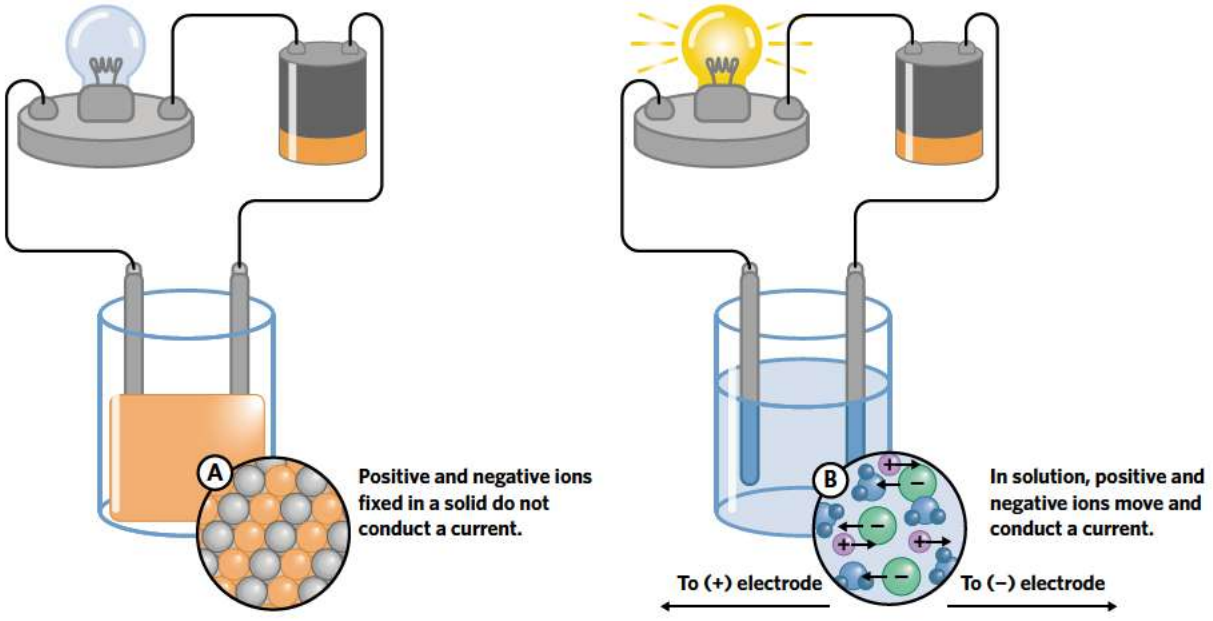


Figure 2 The ability to conduct electricity of ionic compounds in different physical states.

Solubility

As mentioned above, ionic compounds are able to conduct electricity when dissolved in water. However, not all ionic compounds are able to dissolve in water. Those that are able to dissolve in water are referred to as being water **soluble**, whereas ionic compounds that are unable to dissolve in water are referred to as being water **insoluble**. The capacity of different ionic compounds to dissolve in water will be further explored in chapter 7.

Uses of ionic compounds 1.1.12.1

OVERVIEW

The properties of ionic compounds allow them to be widely used for different purposes.

THEORY DETAILS

Due to the fact that ionic compounds are hard, they are commonly used as building materials as well as important structures of the human body as shown in table 1.

Table 1 Applications of ionic compounds due to their hardness.

Ionic compound	Applications
Calcium sulfate ($CaSO_4$)	linings of the walls and house ceilings
Limestone ($CaCO_3$)	building stone
Sandstone (93–93% SiO_2)	building stone
Calcium phosphate	human bone and teeth

Moreover, ionic compounds' high melting points and electrical conductivity in molten and aqueous state allow them to be used widely in many industries as shown in tables 2 and 3, respectively.

Table 2 Applications of ionic compounds due to their high melting points.

Ionic compound	Applications
Magnesium oxide (MgO)	lining surface of furnaces
Lithium fluoride (LiF)	components of engines

Table 3 Applications of ionic compounds due to their electrical conductivity in molten or aqueous state.

Ionic compound	Applications
Ammonium chloride (NH_4Cl)	electrolyte which allows a current flow in cell batteries
Potassium hydroxide (KOH)	electrolyte which allows a current flow in watches and calculators

Other common uses of ionic compounds

- Barium chloride is used to make fireworks.
- Calcium chloride is used to lower the freezing point of water in the winter.
- Sodium chloride is used to make food preservative.
- Cobalt chloride is used to make moisture test papers which change colours when they absorb water.

Theory summary

- Ionic compounds are hard and brittle.
- Ionic compounds have high melting points.
- The ability of ionic compounds to conduct electricity depends on their physical states.
- Some ionic compounds can dissolve in water, whereas others cannot.
- Properties of ionic compounds can be explained using the nature of ionic bonds and the structure of ionic bonds.
- Ionic compounds' properties allow them to be widely used.

4B QUESTIONS

Theory review questions

Question 1

Properties of ionic compounds are determined by the nature of ionic bonds and the structure of 3D ionic crystal lattices.

A True

B False

Question 2

Ionic compounds are hard because

A repulsive force between positive and negative ions are strong.

B ionic bonds between cations and anions are strong.



Question 3

Which of the following best explains the brittleness of an ionic compound?

A When a strong force is exerted on an ionic compound, repulsive forces between like-charged ions will shatter and break the lattice.

B When a strong force is exerted on an ionic compound, repulsive forces between delocalised ions separate the lattice.

Question 4

The majority of ionic compounds have _____ melting points due to the _____ ionic bonds between _____ and _____.

A low, weak, cations, cations

B high, strong, cations, anions

Question 5

Which of the following is true about the electrical conductivity of ionic compounds in solid state?

A Ionic compounds are able to conduct electricity when in solid state due to the movement of delocalised electrons within ionic lattices.

B Ionic compounds are not able to conduct electricity when in solid state because cations and anions are fixed in ionic lattices.

Question 6

In aqueous state, ionic compounds _____ conduct electricity because _____.

A cannot, cations and anions cannot move freely

B can, cations and anions can move freely

Question 7

Which property of limestone ($CaCO_3$) and sandstone (93-93% SiO_2) allows them to be commonly used as building stones?

A Hardness

B High electrical conductivity

Question 8

Magnesium oxide (MgO) is used to make the lining surface of furnaces due to its

A brittleness.

B high melting point.

Question 9

Batteries sometimes contain an electrolyte made of KOH. Electrolytes allow a current to flow. Ionic compounds such as KOH are used for this purpose due to their

A high melting point.

B high electrical conductivity

Question 10

Which property allows an ionic compound to be used to make engine components?

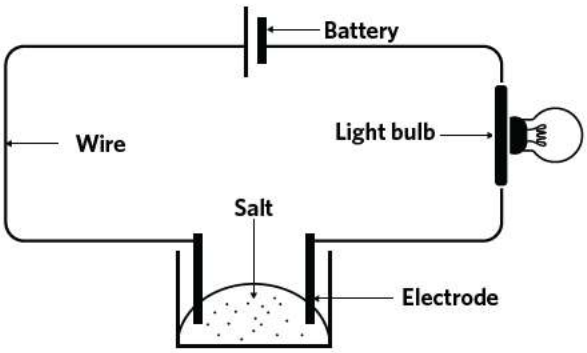
A Brittleness

B High melting point

Deconstructed questions

Use the following information to answer questions 11–13

A group of students, James, Nurit and Grace go camping in the Flinders Ranges in South Australia. After staying there for three days, all of their flashlights have run out of power. Luckily, Nurit brought a battery, a big box of table salt, a bulb, two electrodes and some wires. James uses the solid table salt to set up a simple electrical circuit as shown. However, the light does not shine.



Question 11 (1 MARK)

Table salt is primarily composed of

A sodium oxide.
B sodium chloride.
C magnesium hydroxide.
D magnesium oxide.

Question 12 (1 MARK)

What is the key reason why the bulb does not light?

A There is not enough salt in the box.
B Table salt never conducts electricity.
C The salt is in solid form where all ions are strongly held to each other so there is no movement of ions.
D The salt is in solid form where plenty of ions are moving, which results in big lumps of ions.

Question 13 (3 MARKS)

Suggest one change that can be made to allow the bulb to shine. Justify your answer.

Exam-style questions

Within lesson

Question 14 (3 MARKS)

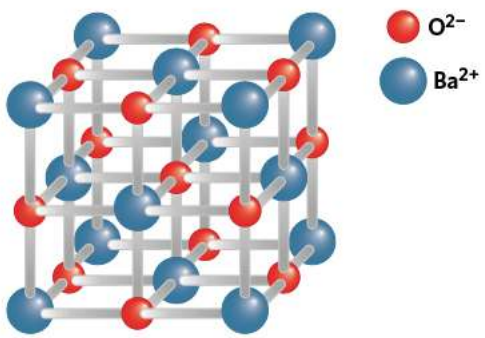
Carrie bought limestone, which is made of calcium carbonate and tested its properties. She used a hammer to break the limestone. In the first three attempts, it was not broken. However, after she hammered it using all of her strength, the limestone was broken into pieces.

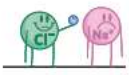
Explain why the limestone was not broken in the first three attempts but broke when Carrie applied a strong force.

Question 15 (3 MARKS)

Barium oxide (BaO) whose lattice structure is given has a melting point of 2000°C.

A Chemistry student explained that barium oxide has a high melting point because of the strong repulsive forces between Ba^{2+} cations. Is the student's explanation correct? Explain your answer.





Question 16 (3 MARKS)

The given table shows the ion content of tap water.

Ion	Percentage (%)
K^+	0.08
Na^+	1.07
Mg^{2+}	1.2
Ca^{2+}	1.4
Cl^-	0.77
SO_4^{2-}	0.7
HCO_3^-	1.8

Explain why people can be electrocuted if a hair dryer is dropped into a bathtub filled with tap water.

Multiple lessons

Question 17 (3 MARKS)

A sodium lattice has a high value of electrical conductivity, while a sodium chloride lattice cannot conduct electricity. Explain why this is the case with reference to the structure of each lattice.

Question 18 (4 MARKS)

Lithium fluoride is widely used to make components of engines due to its high melting point of 848.2°C.

a Use Bohr models to demonstrate the formation of lithium and fluoride ions. (2 MARKS)

b Explain why lithium fluoride has a high melting point. (2 MARKS)

Question 19 (4 MARKS)

Matt wants to produce calcium chloride ($CaCl_2$) from calcium (Ca) and hydrochloric acid (HCl) using the equation:

$Ca_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_{2\,(g)}$

a If Matt uses 10.5 g of calcium, calculate the amount, in mol, of calcium used in the reaction. (1 MARK)

b Why can both calcium chloride solution and solid calcium conduct electricity? (3 MARKS)

Key science skills

Question 20 (6 MARKS)

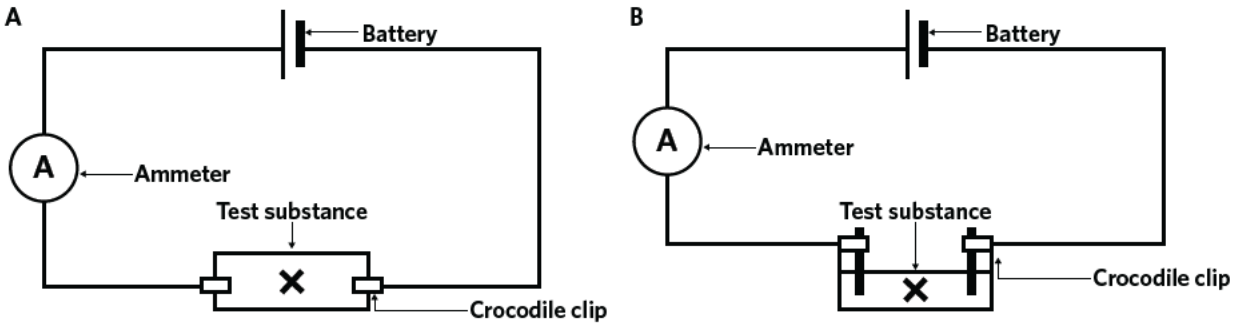
Liam conducted an experiment to investigate the electrical conductivity of sodium chloride (NaCl).

Materials

- 15g of solid sodium chloride (NaCl)
- Distilled water
- 1 × Battery
- Crocodile clips
- 2 × Carbon rods
- Conducting wires
- 1 × Ammeter which is the instrument used to measure the current in an electrical circuit.

Method

1. Use conducting wires to connect the battery and the ammeter as shown in diagram A.
2. Attach each crocodile clip to each end of 5 g of solid NaCl.
3. Leave for 30 seconds.
4. Read the ammeter and record the current value.
5. Remove the solid sodium chloride.
6. Dissolve 4 g of solid NaCl into beaker I containing 500 ml of distilled water.
7. Dissolve 6 g of solid NaCl into beaker II containing 500 ml of distilled water.
8. Use two carbon rods to connect the circuit to the solution in beaker I as shown in diagram B.
9. Leave for 30 seconds.
10. Read the ammeter and record the current value.
11. Repeat steps 8, 9, and 10 for beaker II.



Liam's obtained results are presented in the given table.

Substance	Current (Ampere)
Solid NaCl	0
Solution in beaker I	20
Solution in beaker II	23.5

a Identify the independent and dependent variables. (2 MARKS)

b What can be concluded about the electrical conductivities of solid NaCl, and NaCl solution in beaker I and beaker II? (2 MARKS)

c Explain your conclusion about the electrical conductivities of NaCl solution in beaker I and beaker II. (2 MARKS)

Exam-style question hints

14: Disruptions in the ionic lattice cause the substance to break. **15:** A high melting point indicates that it requires a lot of energy to disrupt the bonds present within a substance. **16:** Moving ions can carry electrical current. **17:** Sodium is a metallic substance. **18a:** Metals donate electrons to non-metals in such a way that both metals and non-metals obtain the octet rule. **18b:** Ionic bonds present in ionic compounds are strong. **19a:** The amount of substance depends on the mass of substance present. **19b:** Free flowing charged particles are required for substances to conduct electricity. **20a:** An independent variable is changed by the experimenter. A dependent variable is the variable being measured or tested. **20b:** A positive current value of ammeter indicates that a substance can conduct electricity. **20c** For ionic compounds in solid state, cations and anions are held by ionic bonds. In aqueous solution, ions can move.

4C WRITING FORMULAS FOR IONIC COMPOUNDS

In this lesson, we will learn how to write formulas for ionic compounds and use these to write full balanced equations involving ionic compounds.

4A Ionic compounds | 4B Properties of ionic compounds | 4C Writing formulas for ionic compounds

Study design dot points

- common properties of ionic compounds (brittleness, hardness, high melting point, difference in electrical conductivity in solid and liquid states) with reference to their formation, nature of ionic bonding and crystal structure including limitations of representations
- precipitation reactions represented by balanced, full and ionic equations, including states

Key knowledge units

Writing ionic formulas	1.1.10.3
Writing balanced equations for the reaction between ionic compounds	1.1.10.4 & 2.1.4.3

Key terms and definitions

Polyatomic ion ion made up of a group of atoms covalently bonded together

Monatomic ions ions made up of only one atom

Double displacement reaction reaction between ionic compounds that results in the exchange of ions to produce new compounds

Lesson links

This lesson builds on:

- 4A Ionic compounds

Cations and anions interact in a ratio that results in a neutral ionic compound.

Writing ionic formulas 1.1.10.3

OVERVIEW

Even though they are made up of charged particles, ionic compounds do not have an overall charge.

THEORY DETAILS

As we discovered in lesson 4A, ionic compounds consist of cations and anions arranged in a lattice structure. Since the lattice structure can go on indefinitely, it would be hard to write a formula that represents the total number of ions involved in the structure of the substance. Unlike molecular formulas, which represent the total number of atoms in a molecule, ionic formulas represent the ratio of ions found in the substance.

When writing ionic formulas, there are few key ideas that we need to keep in mind:

1. The cation is written first, followed by the anion.
2. Ionic compounds are neutral compounds (do not have an overall charge).
3. The charge associated with a **polyatomic ion** represents the overall charge of the ion, not for each single element.
4. Subscripts are used to denote multiples of the same ion.

Consider the ionic compound formed as a result of a reaction between calcium and chlorine. By looking at the periodic table, we can see that calcium is found in Group 2 and chlorine is found in Group 17. Calcium therefore would want to lose its two valence electrons and chlorine would want to gain one electron to have a full valence shell. As a result, calcium donates both of its electrons to become a cation with a charge of +2 and a chlorine, which accepts a donated electron, would become chloride, an anion with a charge of −1 as shown in figure 1.

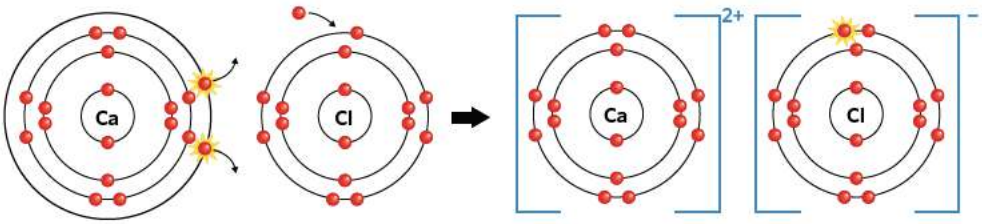


Figure 1 The ionisation of calcium and chlorine atoms.

Since we know the charges of both types of ions, we can write the ionic formula for the compound by balancing the charges, to ensure that the sum of the charges of the cations and anions is equal to zero, as shown in figure 2.

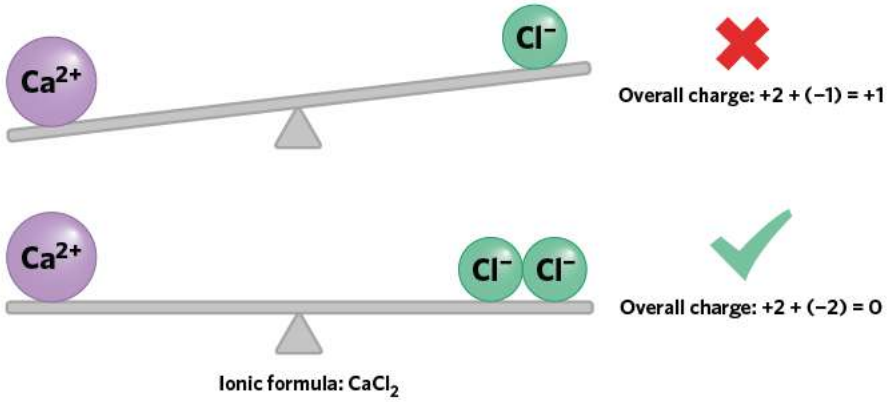


Figure 2 The development of the ionic formula for calcium chloride.

Sometimes, ions exist as polyatomic ions; ions that are made from a group of atoms. The most common polyatomic ions are sulfates (SO_4^{2-}) and carbonates (CO_3^{2-}). When writing ionic formulas involving polyatomic ions, it's important to remember that the ion consists of all of the different atoms as a single unit.



If there are multiple polyatomic ions required, brackets are placed around the polyatomic ion.



Remember to use subscripts to represent multiples of the same monatomic ion or polyatomic ion (with brackets).

Worked example 1

An ionic compound is made up of SO_4^{2-} anions and Na^+ cations. Write the ionic formula for this compound.

What information is presented in the question?

The ions found in the ionic compound.

What is the question asking us to do?

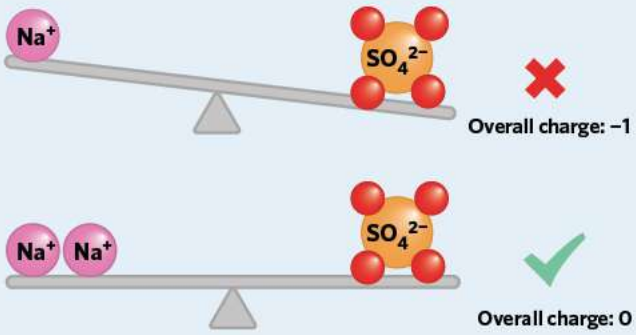
Write an ionic formula for the ionic compound.

What strategy(ies) do we need in order to answer the question?

1. Find the ratio of ions required to form a neutrally charged ionic compound.
2. Used the established ratio to write the ionic formula.

Answer

Based on the information given, we know that there is an imbalance of charge between both ions. We also know that ionic compounds have an overall charge of zero, therefore we would need to balance the charges of each ion so that they cancel out.



As seen, two Na^+ ions are required to balance the charge of a single SO_4^{2-} ion. As such, the ionic formula is Na_2SO_4.



Worked example 2

Write the ionic formula for the ionic compound consisting of CO_3^{2-} and Al^{3+} ions.

What information is presented in the question?

The ions found in the ionic compound.

What is the question asking us to do?

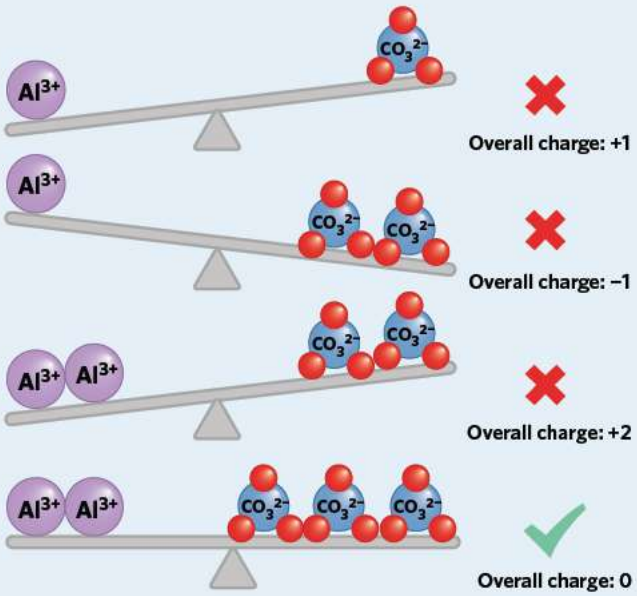
Write an ionic formula for the ionic compound.

What strategy(ies) do we need in order to answer the question?

1. Balance the charge of the ions.
2. Used the established ratio to write the balanced formula.

Answer

Based on the information given, we know that there is an imbalance of charge between both ions. Therefore, we need to find the appropriate ratio of both ions to ensure that the overall charge of the ionic compound is zero.



The ionic formula is $Al_2(CO_3)_3$.

Table 1 shows the charges of common ions.

Table 1 List of common ions and their charges.

Cation			Anions		
Charge	**Name of ion**	**Formula**	**Charge**	**Name of ion**	**Formula**
+1	Ammonium	NH_4^+	−1	Bromide	Br^-
	Copper (I)	Cu^+		Chloride	Cl^-
	Hydrogen	H^+		Hydroxide	OH^-
	Lithium	Li^+		Fluoride	F^-
	Potassium	K^+		Iodide	I^-
	Silver	Ag^+		Nitrate	NO_3^-
	Sodium	Na^+		Hydrogencarbonate	HCO_3^-
+2	Calcium	Ca^{2+}	−2	Carbonate	CO_3^{2-}
	Barium	Ba^{2+}		Sulfate	SO_4^{2-}
	Copper (II)	Cu^{2+}		Sulfite	SO_3^{2-}
	Iron (II)	Fe^{2+}		Sulfide	S^{2-}
	Lead (II)	Pb^{2+}		Oxide	O^{2-}
	Magnesium	Mg^{2+}			
	Nickel (II)	Ni^{2+}			
	Strontium	Sr^{2+}			
	Zinc	Zn^{2+}			
+3	Aluminium	Al^{3+}	−3	Nitride	N^{3-}
	Iron (III)	Fe^{3+}		Phosphate	PO_4^{3-}

If we look closely, we can see that for most of the **monatomic ions**, the charge of the ions can be determined by looking at the periodic table. For example, Na^+, K^+ and Li^+ are metallic elements found in Group 1 on the periodic table. As we know, elements in Group 1 contain one valence electron and therefore these atoms would prefer to lose an electron to have a full outer shell. Consequently, these atoms become ions with a charge of +1. This is not applicable for transition metals as they can form ions with different charges. For these ions, we use Roman numerals to show the charge of the ion. For example, copper (II) has a charge of +2 whereas copper (I) has a charge of +1.



We can use the crisscross method to write ionic formulas. To do so, we can use the number of the charge of each ion as the subscript for the partnering ion. Where necessary, we can then reduce the formula to the lowest ratio.



Naming ionic compounds

There are some simple guidelines when naming ionic compounds:

1 The name consists of two words.
2 The first word is the full name of the metallic element from which the cation comes from.
3 The second word comes from the name of the element from which the anion is derived from, with the ending 'ide' e.g. fluorine → fluoride.
4 Polyatomic ions have unique naming rules that are not the same as monatomic ions.

Worked example 3

Give the name of the ionic compound KBr.

What information is presented in the question?

The ionic formula for the compound.

What is the question asking us to do?

Name the ionic compound with the formula KBr.

What strategy(ies) do we need in order to answer the question?

1. Identify the cation and the anion in the formula.
2. Identify the name of the ionic compound.

Answer

According to the formula, the cation has the elemental symbol K and the anion has the elemental symbol Br.

According to the periodic table, K is the symbol for potassium and Br is the symbol for bromine. To name the ionic compound, we use the full name of the cation, which is potassium and change the name of the element of the anion to bromide. Therefore, the name of this ionic compound is potassium bromide.

By knowing how to read names of ionic compounds, we can write the ionic formula for these compounds.

Worked example 4

Write the formula for copper (I) oxide.

What information is presented in the question?

The name of the ionic compound.

What is the question asking us to do?

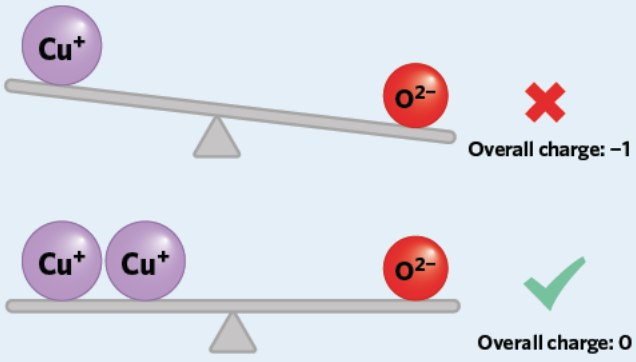
Give the ionic formula for copper (I) oxide.

What strategy(ies) do we need in order to answer the question?

1. Identify the charge of each ion.
2. Find the ratio of ions required for a neutrally charged ionic compound.
3. Use the established ratio to write the balanced formula.

Answer

Based on the table of ions, copper (I) is represented as Cu^+ and oxide is represented as O^{2-}.



The ionic formula for copper (I) oxide is Cu_2O.

Writing balanced equations for the reaction between ionic compounds 1.1.10.4 & 2.1.4.3

OVERVIEW

Some reactions between ionic compounds result in the 'swapping' of ion pairs.

THEORY DETAILS

Ionic compounds can take part in **double displacement reactions**, whereby the ions of two ionic compounds are exchanged to create new substances. Take for example the reaction between sodium chloride (NaCl) and lithium bromide (LiBr) as shown in figure 3.

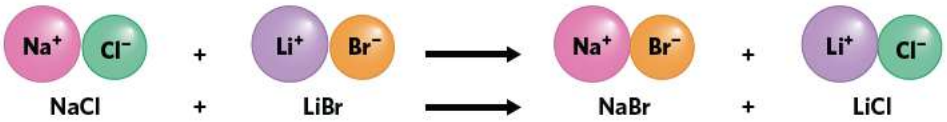
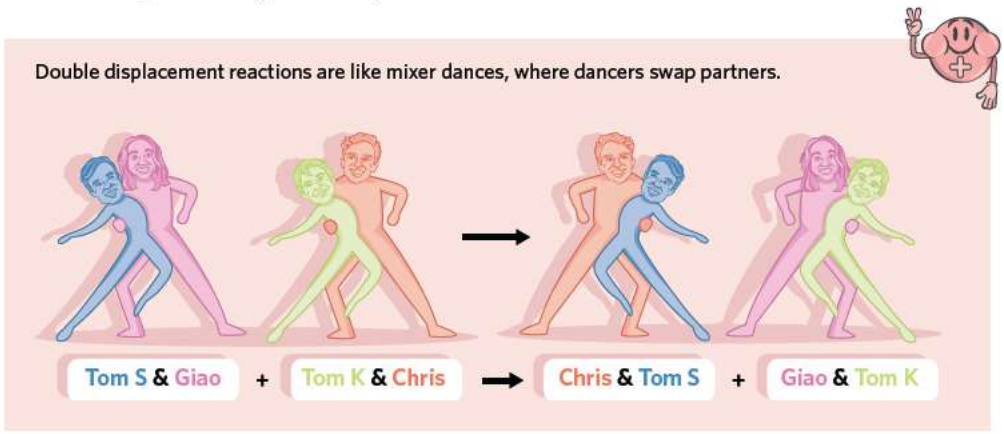


Figure 3 Double displacement reaction between sodium chloride and lithium bromide.

Depending on which perspective we look at it, we can say that there was a swapping of either cations or anions (and not both), resulting in the formation of two new substances. When writing the complete ionic equation for these types of reactions, we need to:

1 Establish the products formed as a result of the exchanging of ions between the reactants.
2 Write the ionic formula for each product.
3 Ensure that both the reactant and product side have the same number of atoms of each element by balancing the full equation.



Worked example 5

Write the balanced ionic equation for a reaction between magnesium chloride ($MgCl_2$) and potassium fluoride (KF).

What information is presented in the question?

The names of each reactant.

What is the question asking us to do?

Write a full balanced ionic equation.

What strategy(ies) do we need in order to answer the question?

1. Establish the products formed as a result of the swapping of ions between both reactants.
2. Identify the ionic formula for each product.
3. Write an unbalanced equation for the reaction.
4. Balance the equation.

Answer

As a result of the reaction, magnesium (originally bonded to chloride) is now bonded to fluoride, and potassium (originally bonded to fluoride) is bonded to chloride.



Now, we need to write the equation for this reaction:

$MgCl_2 + KF \rightarrow MgF_2 + KCl$

Currently, this equation is not balanced, therefore we need to balance the equation to make sure that there are the same number of atoms of each element on both the reactant and product side of the reaction.

$MgCl_2 + 2KF \rightarrow MgF_2 + 2KCl$

This type of reaction will become more important in chapter 7 when we learn about precipitation reactions involving ionic compounds. As the focus of this lesson is to be aware of this type of reaction and to practice our skills, we will not go into great detail about precipitation reactions and will write balanced ionic equations without the use of any states.

Worked example 6

Write the balanced ionic equation for a reaction between sodium nitrate and ($NaNO_3$) and barium iodide (BaI_2).

What information is presented in the question?

The names and formulas of the ionic compounds involved in the reaction.

What is the question asking us to do?

Write a full balanced ionic equation.

What strategy(ies) do we need in order to answer the question?

1. Establish the products formed as a result of the swapping of ions between both reactants.
2. Identify the ionic formula for each product.
3. Write an unbalanced equation for the reaction.
4. Balance the equation.

Answer

As a result of the reaction:



With the new ionic compounds established, we need to make sure to write the correct formulas for each ionic compound:

NaI and $Ba(NO_3)_2$

Now, we need to write the equation for this reaction:

$NaNO_3 + BaI_2 \rightarrow NaI + Ba(NO_3)_2$

Currently, this equation is not balanced, therefore we need to balance the equation to make sure that there are the same number of atoms of each element on both the reactant and product side of the reaction.

$2NaNO_3 + BaI_2 \rightarrow 2NaI + Ba(NO_3)_2$

Theory summary

- Ionic compounds are neutral compounds.
- When establishing the formula for ionic compounds, the total charges of cations and anions present must add up to zero.
- Polyatomic ions are considered as a single unit.
- Ionic compounds can take part in reactions that result in the 'swapping' of ion pairs.
- When writing balanced ionic equations, we need to make sure that the ionic formulas are written accurately and to ensure that there are the same number of atoms of each element on the reactant and product side.



When balancing equations involving polyatomic ions, consider the ion as a single unit, instead of counting the number of atoms in each of the elements found in the ion.

4C QUESTIONS

Theory review questions

Question 1

When writing the formula of an ionic compound, the ion that is first is

A the cation.

B the anion.

Question 2

Ionic compounds can only form between ions of the same charge.

A True

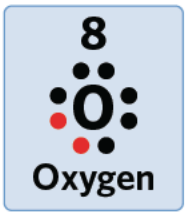
B False



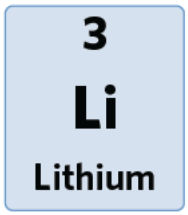
Question 3

Which of the following chemical species is positively charged?

A



B



Question 4

What is the charge of the sulfate ion?

A +2

B −2

Question 5

Which of the following ionic compounds possesses a positively charged polyatomic ion?

A NH_4Cl

B $BaSO_4$

Question 6

What is the name of the ionic compound with formula Ag_2SO_4?

A Sulfate disilver

B Silver sulfate

Question 7

Identify the ionic compound that is written incorrectly.

A $(H)_3PO_4$

B $Al_2(SO_4)_3$

Question 8

Which of the following equations represents a double displacement reaction?

A $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

B $2NaCl + H_2CO_3 \rightarrow Na_2CO_3 + 2HCl$

Question 9

Which of the following represents a balanced, double displacement equation?

A $Na_2SO_4 + AgCl \rightarrow NaCl + Ag_2SO_4$

B $Na_2SO_4 + 2AgCl \rightarrow 2NaCl + Ag_2SO_4$

Question 10

Which of the following represents a double displacement reaction between copper (I) sulfate and iron (III) nitrate?

A $3Cu_2SO_4 + 2Fe(NO_3)_3 \rightarrow 6CuNO_3 + Fe_2(SO_4)_3$

B $CuSO_4 + Fe(NO_3)_2 \rightarrow Cu(NO_3)_2 + FeSO_4$

Deconstructed questions

Use the following information to answer questions 11–13.

Iron exists in a range of different oxidation states which means it can be an ion with a charge of +2, +3, +4 and +6.

Question 11 (1 MARK)

Which of the following correctly represents iron with an oxidation state of +4?

A $Fe(SO_4)_2$
B $Fe(NO_3)_2$
C $FeCl_3$
D $Fe_2(PO_4)_3$

Question 12 (1 MARK)

Which of the following correctly represents iron with an oxidation state of +6?

A $Fe(SO_4)_2$
B $Fe(NO_3)_2$
C $FeCl_3$
D $Fe(PO_4)_2$

Question 13 (2 MARKS)

Write out the balanced double displacement reaction between iron (VI) phosphate and iron (IV) sulfate.
States are not required.

Exam-style questions

Within lesson

Question 14 (10 MARKS)

Complete the following table.

	Name	Formula
a	Silver	
b		Cu^{2+}
c	Hydroxide	
d		F^-
e	Nitrate	

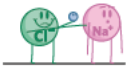
	Name	Formula
f		NH_4^+
g	Sulfate	
h		Fe^{3+}
i	Phosphate	
j		SO_3^{2-}

Question 15 (10 MARKS)

Write out the ionic formula and name the ionic compounds created from the following cation and anion pair.

		Name	Formula
a	Na^+ and NO_3^-		
b	Silver ion and sulfate ion		
c	NH_4^+ and F^-		
d	Hydrogen carbonate ion and strontium ion		
e	K^+ and O^{2-}		

		Name	Formula
f	Calcium ion and carbonate ion		
g	Mg^{2+} and SO_3^{2-}		
h	Hydroxide ion and aluminium ion		
i	Fe^{3+} and S^{2-}		
j	Phosphate ion and Zn^{2+}		



Question 16 (10 MARKS)

Write the balanced double displacement equation for the following pairs of ionic compounds. Note that states are not required.

a Copper (I) sulfate and sodium iodide. (2 MARKS)

b NiO and $FeSO_4$. (2 MARKS)

c Iron (III) nitride and ammonium hydrogencarbonate. (2 MARKS)

d $Ba_3(PO_4)_2$ and HF. (2 MARKS)

e Aluminum sulfite and silver nitrate. (2 MARKS)

Multiple lessons

Question 17 (9 MARKS)

Calcium carbonate is a common mineral found in rocks and is the main component of the shells of marine organisms.

a What is the ionic formula for calcium carbonate? Explain what this formula means in terms of the structure of calcium carbonate. (2 MARKS)

b What property of ionic compounds makes them suitable to be used as shells in marine organisms? Justify your response. (2 MARKS)

Calcium carbonate can be heated to release calcium oxide and carbon dioxide.

c Write the ionic formula for calcium oxide. (1 MARK)

d This decomposition occurs at temperatures greater than 840°C. Why are very high temperatures needed? Justify your response. (2 MARKS)

e Is the following equation involving calcium carbonate a double displacement reaction? Justify your answer. (2 MARKS)

$CaCO_{3(s)} + CO_{2(g)} + H_2O_{(l)} \rightarrow Ca(HCO_3)_{2(aq)}$

Question 18 (10 MARKS)

The thyroid is a gland in the human body which produces important hormones. It is also the only place in the body where iodine can be absorbed so that it can be converted into important hormones. Consequently, iodine is often added in the form of KI to sodium chloride (table salt) to form iodised salt.

a What is the name of KI? (1 MARK)

b What is the ionic formula of sodium chloride? (1 MARK)

c Write out the double displacement reaction between KI and sodium chloride. (1 MARK)

d Silver (I) iodide, used in photography, is produced from the reaction between KI and silver nitrate. Write out the reaction that produces silver (I) iodide. What type of reaction is it? (3 MARKS)

Another use of KI is to stop the thyroid gland absorbing radioactive iodine isotopes, including ^{131}I, after nuclear disasters like Chernobyl.

e Explain the term isotope with reference to ^{131}I and other isotopes of iodine. What is common across all isotopes and what is different? (2 MARKS)

f Radioactivity is caused by neutron release from the nucleus of an unstable isotope. Explain why ^{131}I is radioactive while other isotopes of iodine are found in common table salt. What happens to the mass number of radioactive isotopes over time? (2 MARKS)

Question 19 (12 MARKS)

Salts of fluorine are often added to the municipal water supply and toothpaste in order to prevent dental cavities.

a Fluorine is toxic but the fluoride ion is safe for human consumption. Explain the difference between a fluorine atom and the fluoride ion with respect to their electron configurations and properties. (2 MARKS)

Fluoride ions are essential for the production of fluorapatite which is found in tooth enamel. The ionic formula is $Ca_5(PO_4)_3F$.

b Name and identify all ions present in fluorapatite and their charges. What is the ratio of these ions in fluorapatite. (4 MARKS)

c What properties of fluorapatite make it suitable to function as tooth enamel? Justify your answer. (2 MARKS)

d What is the molar mass of fluorapatite? (1 MARK)

e If a tooth contains 0.25 mg of fluorapatite, how many atoms of oxygen are present? (3 MARKS)

Key science skills

Question 20 (5 MARKS)

Megan is investigating three newly discovered elements, X, Y and Z, which behave as cations in ionic compounds. She knows that the charge of different ions can be determined by the ratio of atoms of each element in an ionic compound. In order to find out the charges, she is able to weigh the mass of each element present and, after a few calculations, determine the unknown charge of the cation.

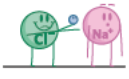
a What type of data is collected? (1 MARK)

b What is the independent and dependent variable? (2 MARKS)

c The masses she recorded in her logbook were based on a scale that wasn't calibrated. What type of error is this and how could her data be corrected? (2 MARKS)

Exam-style question hints

14: Ions have a range of different charges. **15:** Ionic compounds exist in ratios of cations and anions depending on the charge of each. **16:** Double displacement reactions involve the exchange of ions to create new compounds. **17a** Ionic formulas reveal the ratio of anions to cations. **17b:** Ionic compounds have strong electrostatic forces of attraction between ions. **17c** Names of ionic compounds are based on their constituent ions. **17d:** Ionic compounds have strong electrostatic forces of attraction between ions. **17e:** Double displacement reactions involve the exchange of ions to create new compounds. **18a:** Names are based on the ions in an ionic compound. **18b:** Sodium exists as a cation with formula Na^+ and chlorine exists as an anion with a formula Cl^-. **18c:** Double displacement reactions involve the exchange of ions to create new compounds **18d:** Silver nitrate has the ionic formula $AgNO_3$. **18e:** Isotopes of the same element have different mass numbers. **18f:** Mass number is the sum of neutrons and protons in the nucleus of an atom. **19a** Fluoride ions have gained an electron compared to fluorine atoms. **19b:** Ionic formulas reveal the ratio of ions in a compound. **19c:** Ionic compounds are hard and tough. **19d:** Molar mass is the sum of the relative atomic masses of all the atoms present in a compound. **19e:** The number of atoms/molecules present can be calculated using Avogadro's number. **20a** Qualitative data relates to observations while quantitative data relates to numbers. **20b:** The independent variable is deliberately manipulated by the experimenter. **20c:** Errors can either be a one-off or occur throughout the experiment.



EXPERIMENT

ELECTRICAL CONDUCTIVITY OF IONIC COMPOUNDS

Some ionic compounds are able to conduct electricity, however can only do so when in a solution. In this experiment, we will investigate whether the amount of ionic substance in a solution affects the level of electrical conductivity. To do so, we will use an ammeter to measure the flow of electrical current.

Materials

(per group)

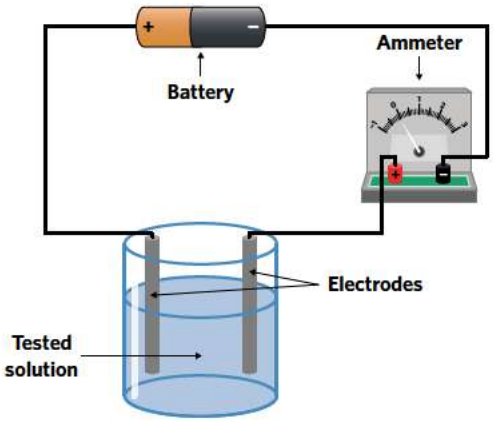
- Solid sodium chloride (NaCl)
- Deionised water
- 1 × Ammeter
- 3 × 250 mL beakers
- 3 × Alligator clips
- 2 × Carbon rods
- 1 × 1.5 V battery
- Paper towel

Method

1 Label the three beakers A, B and C.

2 Make three NaCl solutions based on the table provided.

Beaker	Mass of NaCl (g)	Volume of water (mL)
A	10	100
B	20	100
C	30	100



3 Use an alligator clip to connect the negative terminal of the battery to the negative probe of the ammeter.

4 Using another alligator clip, connect the positive probe of the ammeter to a carbon rod.

5 Use the remaining alligator clip to connect the positive terminal of the battery to the second carbon rod.

6 Place both carbon rods into beaker A, making sure that they do not touch.

7 Record the value shown by the ammeter.

8 Clean and dry the carbon rods and repeat steps 6–7 for the remaining beakers.

Results

Beaker	Mass of NaCl (g)/100 mL of deionised water	Ammeter reading (amps)
A	10	
B	20	
C	30	

QUESTIONS

Question 1 (4 MARKS)

a Identify the dependent variable in the experiment. (1 MARK)

b Identify the general trend shown by your data. (1 MARK)

c Thinking that there was no difference, a student used tap water instead of distilled water in the experiment. Describe the effect that this would have on the data collected. (2 MARKS)

Question 2 (2 MARKS)

Would NaCl be able to conduct electricity in its solid form? Explain. (2 MARKS)

ANSWERS

1 **a** Extent of electrical conductivity (measured in amps).

b As more NaCl is dissolved in the solution, there is an increase in electrical conductivity.

c [Tap water contains traces of different types of ions.[1]][The presence of ions in water can influence the electrical conductivity of the NaCl solution sample, causing an ammeter reading that is higher than the true value.[2]][As a result, this would affect the accuracy of the results.[3]]

✓ ✗ I have identified the presence of ions in water.[1]

✓ ✗ I have described the effect of ions found in water on the electrical conductivity of the NaCl solution.[2]

✓ ✗ I have identified the overall effect on the experimental results.[3]

2 [Free flowing charged particles are required to conduct electricity.[1]][When in solid form, the sodium and chloride ions in NaCl are unable to move.[2]][Therefore, the solid sodium chloride would not be able to conduct electricity.[3]]

✓ ✗ I have identified the characteristic that is necessary to conduct electricity.[1]

✓ ✗ I have identified the key characteristic of solid NaCl related to electrical conductivity.[2]

✓ ✗ I have linked my answer to the question.[3]



CHAPTER 4 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Question 1 (1 MARK)

Most ionic compounds are formed from

A metal and non-metal cations.
B a non-metal cation and a metal anion.
C metal anions.
D a metal cation and a non-metal anion.

Question 2 (1 MARK)

An ionic compound

A is a three-dimensional structure formed from a single cation and anion.
B involves single ions forming multiple ionic bonds with multiple other ions.
C contains strong covalent bonds between ions.
D is a two-dimensional structure formed from a single cation and anion.

Question 3 (1 MARK)

The forces between chloride ions will be _____ forces, and the forces between sodium and chloride ions will be _____ forces.

A repulsive, electrostatic
B electrical, repellant
C electrostatic, repulsive
D electromagnetic, hydrostatic

Question 4 (1 MARK)

An ionic compound left to cool for a longer period of time will

A have smaller crystals.
B have a higher melting point.
C have larger crystals.
D be in a molten state.

Question 5 (1 MARK)

Ionic compounds are

A insulators of electricity.
B effective conductors of electricity.
C effective conductors of electricity in a molten or dissolved state.
D effective conductors of electricity in a solid state.

Question 6 (1 MARK)

Which of the following best describes solid ionic compounds?

A Soft, malleable
B Hard, brittle
C Soft, flexible
D Hard, malleable

Question 7 (1 MARK)

Lithium fluoride (LiF) is used to produce components of engines due to its

A high melting point.
B hardness.
C low melting point.
D tensile strength.

Question 8 (1 MARK)

Which of the following is the ionic formula for an ionic compound made from magnesium and fluoride ions?

A Mg_2F
B FMg_2
C MgF_2
D F_2Mg

Question 9 (1 MARK)

Which of the following is the correct ionic formula for iron (III) hydroxide?

A $FeOH$
B Fe_3OH
C $FeOH_3$
D $Fe(OH)_3$

Question 10 (1 MARK)

Which of the following represents the balanced equation of the reaction between calcium carbonate ($CaCO_3$) and aluminium chloride ($AlCl_3$)?

A $3CaCO_3 + 2AlCl_3 \rightarrow Ca_3Al_2 + 3Cl_2CO_3$
B $CaCO_3 + AlCl_3 \rightarrow CaCl_3 + AlCO_3$
C $3CaCO_3 + 2AlCl_3 \rightarrow 3CaCl_2 + Al_2(CO_3)_3$
D $2CaCO_3 + 3AlCl_3 \rightarrow Ca_2Cl_9 + Al_3(CO_3)_2$

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (10 MARKS)

Ionic compounds are found abundantly in the natural world, and also serve numerous purposes in human activities. An example of an abundant ionic compound is sodium chloride, which is found in high concentrations in seawater, and is also used as table salt.

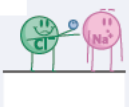
a Write the ionic formula of sodium chloride. (1 MARK)

b The current Australian dietary guidelines state that no more than six grams of salt should be consumed daily due to the negative effects of excessive sodium on blood pressure.

i A customer at a restaurant finds 0.13 mol of table salt in their meal. Determine whether this is a healthy amount of salt, showing your calculations. (2 MARKS)

ii The customer hypothesises that by heating up food containing salt at a moderate temperature, the salt will degrade and thus not have the same health consequences. Comment on the customer's hypothesis, referring to the properties of ionic compounds. (2 MARKS)

c The customer obtains a block of pure sodium chloride and uses a hammer to hit the salt block as hard as they can. Describe what would most likely be observed, with reference to the structure and properties of ionic compounds. (2 MARKS)



d Complete the table which demonstrates the different ways in which ionic compounds can be used. (3 MARKS)

Use	Ionic compound	Properties for this use
Lining surface of engines	Magnesium oxide (MgO)	
Human bones and teeth	Calcium phosphate ($Ca_3(PO_4)_2$)	
Electrolyte enabling current flow in cell batteries	Ammonium chloride (NH_4Cl)	

Question 12 (13 MARKS)

When determining the equation for reactions between ionic compounds, it is important to take the charges of compounds into account.

a Identify the name of reactions between ionic compounds and describe what takes place. (2 MARKS)

b Provide the ionic formula and name of the ionic compound consisting of Ba^{2+} and PO_4^{3-} ions. (2 MARKS)

c Write the balanced equation for the reactions between the following ionic compounds. Note: States are not required.

i Calcium bromide ($CaBr_2$) and copper (I) sulfite (Cu_2SO_3). (1 MARK)

ii Silver carbonate and potassium fluoride. (1 MARK)

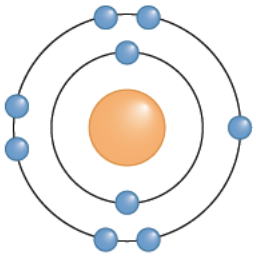
iii Copper (II) nitride and sodium sulfide. (1 MARK)

d Provide the names of the products of the reactions in q12ci–iii. (3 MARKS)

e A new atom, J, was discovered, however scientists were unsure about the properties of the atom. In a reaction with the non-metal oxygen, it was discovered that the substance produced had a formula of J_2O.

i Would J be considered as being a metal or non-metal? (1 MARK)

ii If J was reacted with the atom shown, identify the formula of the compound formed. (1 MARK)



iii Which group on the periodic table should atom 'J' be added to? (1 MARK)

Question 13 (7 MARKS)

Calcium fluoride is sometimes used in the manufacture of windows.

a What does calcium fluoride's formula suggest about its structure? (2 MARKS)

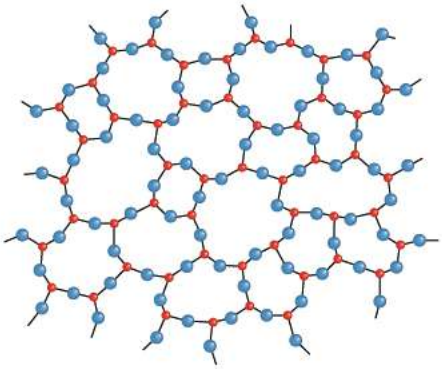
b Win and Co, a manufacturing company, wanted to develop a new type of window that is highly scratch resistant and strong. A general overview of the production process is shown.



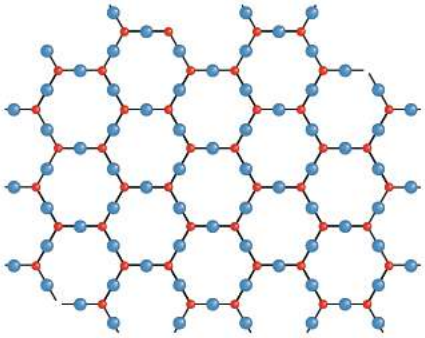
i Which of the stages of production would have the greatest impact on the strength of the window? Explain. (2 MARKS)

ii To investigate the structure of the windows made, engineers at the company took high resolution imaging of some sample windows immediately after stage 2 of the overall process.

Sample 1



Sample 2



Compare the conditions of sample 1 and 2. (2 MARKS)

c It is believed that the smaller the crystals formed, the stronger the material. Based on the results in part b, identify the sample whose condition most favours the production of a strong window. (1 MARK)

KEY SCIENCE SKILLS QUESTIONS (7 MARKS)

Question 14 (7 MARKS)

A student who has recently learned about ionic compounds has prepared an experiment to compare the properties of different ionic compounds. The experiment involves exposing a series of different samples of compounds to a flame for a set amount of time and observing the interactions of each with the flame.

a Identify the independent and dependent variables in this experiment. (2 MARKS)

b Identify at least two controlled variables in this experiment. (2 MARKS)

c The student gave two other students a copy of the method and each of them carried out the experiment using the same samples, all accounting for control variables. The results are shown in the table:

	Student 1	**Student 2**	**Student 3**
Magnesium hydroxide ($Mg(OH)_2$)	Melts	No change	Melts
Sodium chloride (NaCl)	No change	Hardens	Melts
Lead (II) Sulfide (PbS)	Hardens	Hardens	Melts

i Is this experimental method reliable? Explain your answer. (2 MARKS)

ii It was found that each of the students had a different standard metric that they were using to determine the change the sample had undergone. Identify what type of error this is. (1 MARK)

UNIT 1

AOS 2

How can the versatility of non-metals be explained?

In this area of study students explore a wide range of substances and materials made from non-metals including molecular substances, covalent lattices, carbon nanomaterials, organic compounds and polymers.

Students investigate the relationship between the electronic configurations of non-metallic atoms and the resultant structures and properties of a range of molecular substances and covalent lattices. They compare how the structures of these non-metallic substances are represented and analyse the limitations of these representations. Students study a variety of organic compounds and how they are grouped into distinct chemical families. They apply rules of systematic nomenclature to each of these chemical families. Students investigate useful materials that are made from non-metals, and relate their properties and uses to their structures. They explore the modification of polymers and the use of carbon-based nanoparticles for specific applications.

Students apply quantitative concepts to molecular compounds, including mole concept and percentage composition by mass, and determine the empirical and molecular formulas of given compounds.

Outcome 2

On completion of this unit the student should be able to investigate and explain the properties of carbon lattices and molecular substances with reference to their structures and bonding, use systematic nomenclature to name organic compounds, and explain how polymers can be designed for a purpose.

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UNIT 1 AOS 2, CHAPTER 5

Molecular theory

05

5A Covalent bonding

5B Properties of covalent compounds

5C Intermolecular bonding

Key knowledge

- representations of molecular substances (electron dot formulas, structural formulas, valence structures, ball-and-stick models, space-filling models) including limitations of representations
- shapes of molecules and an explanation of their polar or non-polar character with reference to the electronegativities of their atoms and electron-pair repulsion theory
- explanation of properties of molecular substances (including low melting point and boiling point, softness, and non-conduction of electricity) with reference to their structure, intramolecular bonding and intermolecular forces
- the relative strengths of bonds (covalent bonding, dispersion forces, dipole-dipole attraction and hydrogen bonding) and evidence and factors that determine bond strength including explanations for the floating of ice and expansion of water at higher temperatures

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5A COVALENT BONDING

In this lesson, we'll look at the bonds which hold molecules together and how these molecules are represented.

5A Covalent bonding | 5B Properties of covalent compounds | 5C Intermolecular bonding

Study design dot points

- the relative strengths of bonds (covalent bonding, dispersion forces, dipole-dipole attraction and hydrogen bonding) and evidence and factors that determine bond strength including explanations for the floating of ice and expansion of water at higher temperatures
- representations of molecular substances (electron dot formulas, structural formulas, valence structures, ball-and-stick models, space-filling models) including limitations of representations

Key knowledge units

Covalent bonding	1.2.4.1
Representing molecules	1.2.1.1

Key terms and definitions

Molecule two or more atoms chemically bonded together

Intramolecular bonds bonding between atoms within a molecule

Covalent bonds bonding resulting from the sharing of electrons

Octet rule general principle which states that atoms like to have 8 electrons in their valence shell for stability

Single covalent bond bond formed by the sharing of two electrons/one pair of electrons

Diatomic molecule molecule consisting of only two atoms (they may be different elements)

Double covalent bond bond formed by the sharing of four electrons/two pairs of electrons (two covalent bonds)

Triple covalent bond bond formed by the sharing of six electrons/three pairs of electrons (three covalent bonds)

Organic chemistry chemistry of carbon-containing compounds

Electron configuration arrangement of electrons in shells and/or subshells

Valence electrons electrons in the outermost shell of an atom

Electron dot formula representation of a molecule where valence electrons are shown as dots around the chemical symbol

Unpaired electron electron in an orbital by itself

Lone pair of electrons pairs of valence electrons not shared with another atom in a covalent bond

Inner shell electrons electrons not in the valence shell

Valence structure representation of a molecule where chemical symbols are written (as well as valence electrons), joined by single, double or triple lines, corresponding to the type of bond

Ball-and-stick model representation of a molecule demonstrating the three-dimensional arrangement of atoms

Space-filling model representation of a molecule demonstrating the three-dimensional arrangement, and the respective shape and size of atoms

Lesson links

This lesson builds on:

▶ 3A Metals

Metallic bonding is another type of intramolecular bonding.

▶ 4A Ionic compounds

Ionic bonding is another type of intramolecular bonding.

Covalent bonding 1.2.4.1

OVERVIEW

Covalent bonding is responsible for holding together non-metallic atoms.

THEORY DETAILS

A **molecule** is a group of atoms chemically bonded together. In this lesson, we will explore one type of bond that holds these molecules together. Bonds between atoms of the same molecule are called **intramolecular bonds**.

One of the types of intramolecular bonding for non-metallic substances is **covalent bonding** which involves the sharing of electrons, as opposed to ionic bonding, which involves the complete transfer of electrons. This difference is illustrated in figure 1.

Due to their atomic structure, non-metallic atoms like to share electrons in order to complete their valence shell. Most of the time, this is due to the **octet rule** which describes that atoms are most stable when they have eight electrons in their valence shell. The electrons that are shared between the atoms are attracted to the positive nucleus of both atoms. This force of attractions holds the atoms together.

When a bond is formed by the sharing of two electrons, it is referred to as a **single covalent bond**. Hydrogen gas (H_2) is an example of a molecule with a single covalent bond. In this case, each hydrogen atom shares its single electron with the other hydrogen atom, resulting in a intramolecular bond as shown in figure 2.

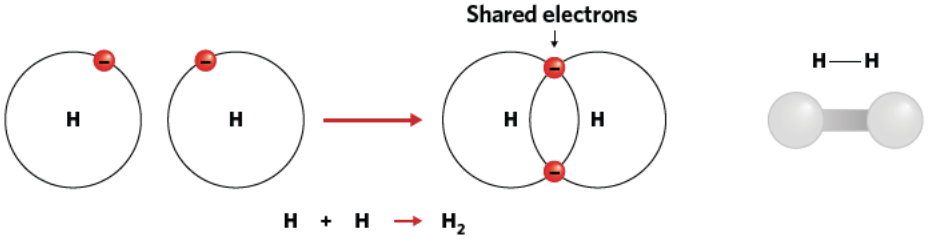


Figure 2 Formation of $H_{2(g)}$ from two hydrogen atoms.

In this case, since the valence shell of a hydrogen atom can only hold two electrons, the sharing of the two electrons completes the valence shell of both atoms, forming a covalent bond. Hydrogen gas (H_2) is also known as a **diatomic molecule** as it consists of two atoms covalently bonded together.

However, it is also possible for an atom to share more than one electron. For example oxygen, which has six valence electrons, needs two electrons to complete its outer shell. Therefore, since each oxygen atom can share two electrons and each oxygen atom wants another two electrons, a **double covalent bond** can form by the sharing of a total of four electrons (or two pairs of electrons) as shown in figure 3.

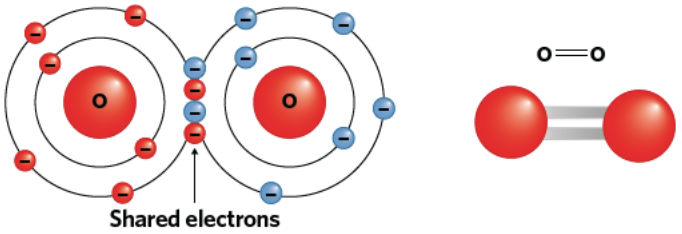


Figure 3 Double covalent bond in $O_{2(g)}$ and each shared pair of electrons represents one covalent bond.

Some atoms have the ability to share three electrons. Nitrogen is one example as it has five valence electrons. Therefore, according to the octet rule, it requires three more electrons to fill its valence shell. When two nitrogen atoms bond together, each atom shares three electrons with the other, forming a **triple covalent bond** as shown in figure 4.

Another important thing to note is that the greater the number of electrons shared in the covalent bond, the stronger the bond. This means that in general, a triple covalent bond is stronger than a double covalent bond which is stronger than a single covalent bond. All of these examples have been between two atoms of the same type. However, covalent bonds can exist between atoms of different elements. Carbon, which is a very common element, has four valence electrons.



Intramolecular bonding is bonding between atoms (as opposed to intermolecular bonding which will be covered in a later lesson).

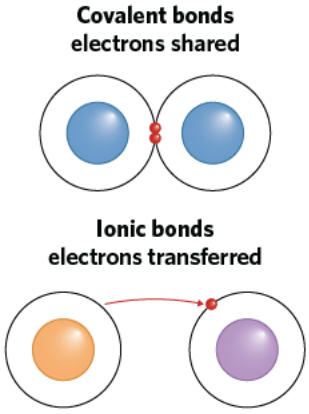


Figure 1 Difference between covalent bonding and ionic bonding.



Some exceptions to the octet rule include hydrogen, sulphur, beryllium and boron atoms.



Whenever an atom shares an electron with another atom, a covalent bond is formed.



Whenever an atom shares two electrons with another atom which shares two electrons as well, a double covalent bond is formed.

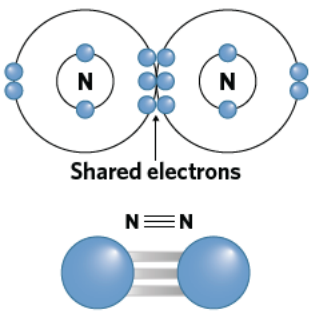


Figure 4 Triple covalent bond in $N_{2(g)}$, each shared pair of electrons represents one covalent bond.



This means that it is very good at forming covalent bonds with other atoms because it wants to complete its outer shell. In fact, it can covalently bond with up to four other atoms and this is one explanation for the diversity of carbon-containing compounds, a field called **organic chemistry** which will be explored later.



Whenever an atom shares three electrons with another atom which shares three electrons as well, a triple covalent bond is formed.

Naming covalent compounds

The naming or nomenclature of covalent compounds can help us to write the molecular formula for the compound and vice versa. The naming of covalent compounds follows some very simple rules. For the purpose of this study design, we will look at the naming of covalent compounds consisting of two elements. Take for example the molecule NO_2.

NO_2

Nitrogen dioxide

The first part of the compound name is taken from the full name of the first element (from left to right) in the formula, followed by a slightly adjusted version of the second element found in the compound. This is generally written as the 'anion' version (negatively charged ion) of the element. For example, for oxygen, this would be 'oxide'. It's very important to note that this does not mean that the compound is ionic, it is purely a way of naming the compound. At times, names can include prefixes that show the number of atoms present. For the molecule NO_2, the prefix *di-* is used meaning two as there are two oxygen atoms bonded to the nitrogen atom. Table 1 shows the common prefixes used.

Table 1 Prefixes representing different numbers.

Prefix	Number represented
mono-	1
di-	2
tri-	3
tetra-	4
pent-	5

Representing molecules 1.2.1.1

OVERVIEW

There are many different ways to represent molecules and each one has their advantages and disadvantages.

THEORY DETAILS

As we learned in lesson 2A, atoms have **electron configurations**, with the electrons in the outer shell of an atom named **valence electrons**. These valence electrons are important for bonding and the shape of molecules.

Due to the importance of these valence electrons, **electron dot formulas** (also called Lewis structures or electron dot diagrams) illustrate the number of valence electrons that an atom has with dots representing electrons as shown in figure 5.

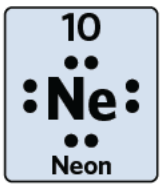


Figure 5 Electron dot formula of neon.

In this case, the **inner shell electrons** of neon (there are two of them) are not shown because they do not contribute to the bonding or shape of neon. Instead, only the valence electrons are shown. If an atom has four or less valence electrons, they are represented equally spaced around the atomic symbol and if there are more than four electrons, they are represented in pairs as shown in figure 6.

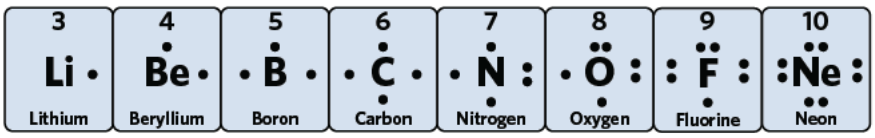


Figure 6 Electron dot formulas of lithium to neon.

Having seen some of these common atoms by themselves, we can start to see the electron dot formulas of covalent molecules, which are atoms covalently bonded together through the sharing of electrons. The advantage of using an electron dot formula is that it is able to tell us which atoms have **unpaired electrons**. Unpaired electrons are special because these electrons are able to participate in bonding while paired electrons are not able to be involved in bonding. This is why some atoms like neon, which have no unpaired electrons and a full valence shell, are unreactive and very rarely undergo any form of bonding. Electron dot formulas can also be used to represent molecules and the bonds between atoms. For example, when two atoms of chlorine are bonded together to form Cl_2 (chlorine gas), the unpaired electrons are shared allowing a single covalent bond to form as shown in figure 7. This configuration is favourable due to the octet rule which states that for the majority of molecules the most stable configurations involve eight electrons in the valence shell.

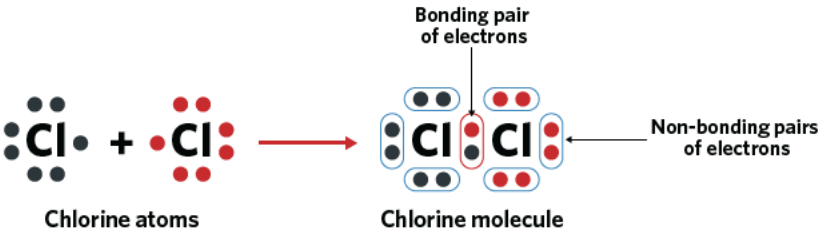


Figure 7 Electron dot formula of Cl_2.

In this case, since both chlorine atoms now have 8 electrons in their outer shell, they are more energetically stable bonded together. In figure 7, the bonding pair of electrons and the non-bonding pairs of electrons are highlighted. Non-bonding pairs of electrons are also called **lone pairs of electrons** as they are not shared with another atom in a covalent bond. It is also important to understand that electron dot formulas are relatively simple and cannot explain all properties of molecules. Furthermore, for larger molecules, the number of electrons is greatly increased meaning that electron dot diagrams can start to appear complicated.



Number of valence electrons < four = equally spaced, number of valence electrons > four = paired.

Worked example 1

Draw an electron dot diagram for the molecule PF_3.

What information is presented in the question?

The molecule is PF_3.

What is the question asking us to do?

Draw an electron dot diagram.

What strategy(ies) do we need in order to answer the question?

1. Use the periodic table to determine the number of electrons in each atom of P (phosphorus) and F (fluorine).
2. Determine the number of valence electrons in each atom by finding the electron configuration.
3. Place four of the valence electrons of each atom around the atomic symbol, and then start pairing the electrons up.
4. Draw an electron dot diagram taking into account that a single bond involves two shared electrons.

Answer

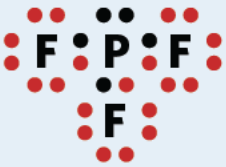
Phosphorus has an atomic number of 15 which means in a neutrally charged atom of phosphorus there are 15 electrons. Using our rules of electron configurations, there will be two electrons in the first shell, eight electrons in the second shell and the remaining five electrons in the third shell. This gives an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^3$. This means that the number of valence electrons is five.



Fluorine has an atomic number of 9 which means in a neutrally charged atom there are nine electrons. Using our knowledge of electronic configuration, there will be two electrons in the first shell and seven electrons in the second shell. This gives an electron configuration of $1s^2 2s^2 2p^5$. Therefore, fluorine has seven valence electrons.



Since the three fluorine atoms would like a complete octet of electrons, they each form a single bond with the phosphorus atom. Therefore, the electron dot diagram is as follows:



Electron dot diagrams can also be used to represent double and triple covalent bonds. As mentioned, double bonds are a result of the sharing of four electrons. This is represented by both oxygen atoms in figure 8, both sharing their two unpaired electrons. The same can be said for triple bonds. The electron dot diagram of nitrogen in figure 8 shows that both nitrogen atoms share their three unpaired electrons, resulting in a triple bond involving six shared electrons.





Figure 8 Electron dot diagrams of oxygen (double bond) and nitrogen (triple bond).

Worked example 2

Draw an electron dot diagram for the molecule carbon dioxide (CO_2).

What information is presented in the question?

The molecule is CO_2.

What is the question asking us to do?

Draw an electron dot diagram.

What strategy(ies) do we need in order to answer the question?

1. Use the periodic table to determine the number of electrons in each atom of C (carbon) and O (oxygen).
2. Determine the number of valence electrons in each atom by finding the electron configuration.
3. Place four of the valence electrons of each atom around the atomic symbol, and then start pairing the electrons up.
4. Draw an electron dot diagram taking into account that a double bond involves four shared electrons.

Answer

Carbon has an atomic number of 6 which means in a neutrally charged atom of carbon there are six electrons. Using our rules of electron configurations, there will be two electrons in the first shell and four electrons in the second shell. This gives an electron configuration of $1s^2 2s^2 2p^2$. This means that the number of valence electrons is four. Since there are only four valence electrons, they will be equally spaced around the atomic symbol. Furthermore, carbon is likely to be the central atom in the electron dot diagram because it has the greatest number of unpaired valence electrons.



Oxygen has an atomic number of 8 which means in a neutrally charged atom there are eight electrons. Using our knowledge of electronic configuration, there will be two electrons in the first shell and six electrons in the second shell. This gives an electron configuration of $1s^2 2s^2 2p^4$. Therefore, oxygen has six valence electrons. Since there are six valence electrons, after placing four electrons equally spaced around the atomic symbol, two more electrons will be coupled with another electron.



Since the carbon atom wants four electrons (has four unpaired electrons) and each oxygen atom wants two electrons (has two unpaired electrons), each oxygen can form a double bond with the central carbon atom as shown below.



After arranging this into an electron dot diagram, the final answer is:



A **valence structure** is a different form of the electron dot diagram where the bonding electrons are replaced with a single line. This can be seen in figure 9.

The line represents a shared pair of electrons, which is a single bond. For molecules with double or triple bonds, the single line is replaced with a double or triple line as shown in figure 10.

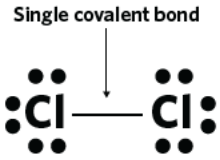


Figure 9 Valence structure of Cl_2.

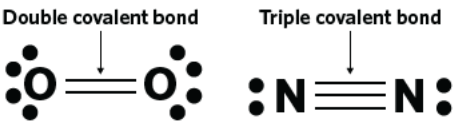


Figure 10 Valence structure of O_2 and N_2.

Structural formulas are one step more simplified than electron dot formulas and valence structures as shown in figure 11.

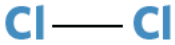


Figure 11 Structural formula of Cl_2.

In this structure, only the bonds between atoms are represented. This provides a cleaner approach to representing molecules but it does not provide information about the valence electrons.



Given the structural formula, it is possible to convert it to an electron dot formula or valence structure by adding valence electrons.

Another approach to representing molecules is displaying the molecule with its three-dimensional arrangement of atoms. There are two types of models for this. The first is the **ball-and-stick model** in figure 12 which shows the atoms in the molecule and their three-dimensional arrangement. In this diagram, the molecule is the organic compound butane C_4H_{10} and the black spheres represent carbon atoms while the white spheres represent hydrogen atoms. The atoms are shown as much smaller than the bonds between the atoms which means that this model does not represent the actual size and relative scale of atoms. Therefore, an alternative to the ball-and-stick model is the **space-filling model** where the atomic radius of atoms is accurate as well as the distance between the nuclei of atoms, as shown in figure 13.

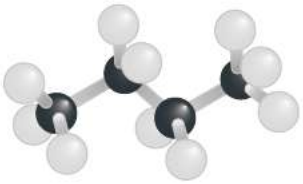


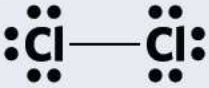
Figure 12 Ball-and-stick model of butane.



Figure 13 Space-filling model of butane.

A summary of the different ways of representing molecules, their advantages and disadvantages is shown in table 2.

Table 2 Comparison of different ways of representing molecules.

Model	Example	Advantages	Disadvantages
Electron dot formula		Clearly shows which electrons are involved in bonding	Does not show the three-dimensional shape of the molecule Is complicated and time-consuming to write
Valence structure		Simplifies the electron dot formula	Does not show the three-dimensional shape of the molecule
Structural formula		Simplifies the valence structure to show just the bonds between molecules	Does not show the three-dimensional shape of the molecule
Ball-and-stick model		Displays the shape of the molecule	Doesn't show the relative sizes of atoms
Space-filling model		Displays the relative sizes of atoms	Does not show bond angles clearly and the different types of bonds

Theory summary

- Covalent bonding is a form of intramolecular bonding caused by the sharing of electrons due to the fact that atoms like to fill their valence shell according to the octet rule.
- Single, double, triple bonds are formed from the sharing of two, four and six electrons respectively.
- Electron dot formulas and valence structures focus on the valence electrons of atoms.
- Structural formulas focus on the bonds between atoms.
- Ball-and-stick and space-filling models focus on the three-dimensional structure and shape of molecules.



5A QUESTIONS

Theory review questions

Question 1

A molecule consists of a

A single atom.

B group of atoms bonded together.

Question 2

Electrons are transferred in a covalent bond.

A True

B False

Question 3

Which of the following depicts a double covalent bond?

A H:H

B 

Question 4

Electron dot formulas can be used to represent

A both individual atoms and molecules.

B only individual atoms.

Question 5

The dots in electron dot formulas represent

A inner shell electrons.

B the electrons in the outer shell of an atom.

Question 6

Valence structures are most similar to

A electron dot formulas.

B ball-and-stick models.

Question 7

The ball-and-stick model is a more accurate three-dimensional representation of molecules than the space-filling model.

A True

B False

Question 8

All atoms obey the octet rule.

A True

B False

Question 9

The central atom in the pictured molecule can have 12 electrons in its valence shell. Does it obey the octet rule?



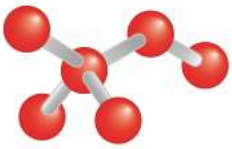
A Yes

B No

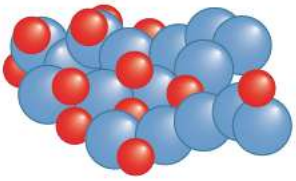
Question 10

Which of the following images depicts the space-filling model?

A



B



Deconstructed questions

Use the following information to answer questions 11–13

Tom would like to draw the electron dot diagram of a molecule called Tilden's reagent, NOCl (nitrosyl chloride). He is unaware of the number of bonds and whether there are double or even triple bonds present.

Question 11 (1 MARK)

How many valence electrons are there in nitrogen, oxygen and chlorine respectively?

A 5, 6, 7

B 6, 7, 5

C 7, 6, 7

D 6, 5, 7

Question 12 (1 MARK)

Some atoms have more valence electrons than others. In order to obey the octet rule, which atom needs to have the most electrons shared with it?

A Nitrogen

B Oxygen

C Chlorine

D All of the above

Question 13 (2 MARKS)

Based on the above information, draw the electron dot diagram of nitrosyl chloride, NOCl.

Exam-style questions

Within lesson

Question 14 (10 MARKS)

For each of the following molecules, draw the electron dot diagram. Remember that sulphur and nitrogen atoms can sometimes disobey the octet rule.

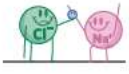
a CH_4 (2 MARKS)

b CO_2 (2 MARKS)

c SO_3 (2 MARKS)

d NO (2 MARKS)

e CO (2 MARKS)



Question 15 (7 MARKS)

Chris, a keen chemist, would like to determine which molecular representation he should use for a molecule which he recently discovered. In order to start the process, he begins to compare the different representations.

a What is one similarity and one difference between electron dot diagrams and valence structure representations? (2 MARKS)

b What is one similarity and one difference between the space-filling model and ball-and-stick model? (2 MARKS)

The molecule he discovered has many large atoms and he wants to be able to visualise its shape in three-dimensions.

c Given this information, from the four previously listed representations, which should Chris use? Justify your answer. (3 MARKS)

Question 16 (4 MARKS)

Oxygen difluoride, OF_2 was first noticed in 1929 and reacts with many metals, forming oxides.

a Draw the electron dot diagram of oxygen difluoride. (1 MARK)

b Circle the lone pairs of electrons. (1 MARK)

c Why does oxygen, when it is involved in a molecule like OF_2, usually have 2 lone pairs of electrons? (2 MARKS)

Question 17 (7 MARKS)

Hydrogen peroxide (H_2O_2) and oxygen gas (O_2) both have an oxygen atom covalently bonded to another oxygen atom.

a Identify the type of oxygen-oxygen covalent bond present in O_2. Explain your answer. (2 MARKS)

b Identify the type of oxygen-oxygen covalent bond present in H_2O_2. Explain your answer. (2 MARKS)

c Hydrogen peroxide undergoes explosive chemical reactions whereas oxygen gas is a stable molecule and crucial for life. With reference to the type of covalent bonds present in both molecules, offer a possible explanation for this dramatic difference in properties. (3 MARKS)

Question 18 (10 MARKS)

Katherine was investigating the properties of different atoms and how the maximum number of covalent bonds that they could form impacted the number of molecules that they could form.

a How many covalent bonds can neon form? Explain your answer. (2 MARK)

b What combination of single, double and triple covalent bonds does oxygen usually form? Explain your answer. (3 MARKS)

c What combination of single, double and triple covalent bonds can carbon form? List all combinations possible where the carbon atom has a full valence shell. Explain your answer. (3 MARKS)

d Using the above as evidence, explain why there are more than 9 000 000 organic compounds (those with carbon present). (2 MARKS)

Multiple lessons

Question 19 (6 MARKS)

Hermione was chosen by her teacher to research the difference between ionic and covalent bonds for a class presentation.

a Describe the difference between ionic and covalent bonds. (2 MARKS)

CCl_4 and NaCl are both molecules with a chlorine atom. One of them is covalently bonded and the other is ionically bonded.

b Which one of CCl_4 or NaCl is covalently bonded and which one is ionically bonded. Give two theoretical reasons why you have reached that conclusion. (2 MARKS)

c Give two ways of determining which one is ionically bonded using experiments. (2 MARKS)

Key science skills

Question 20 (5 MARKS)

Peter and Paul were investigating the type of bonding present in unknown chemical compounds by using a technique called melting point determination. In the experiment, they used a machine called a capillary melting point apparatus which melts a small sample of a chemical compound in a tiny capillary tube. The device is shown.



Image: SUKJAI PHOTO/Shutterstock.com

The results collected after investigating substance A and B using the capillary melting point apparatus are shown below.

	Substance A	Substance B
Melting point (°C)	65	756

a Identify the independent variable. (1 MARK)

b Write an appropriate aim for the experiment. (1 MARK)

c What type of data is collected in the table? (1 MARK)

d Which substance, A or B, is more likely to be covalently bonded? Explain your answer. (2 MARKS)

Exam-style question hints

14: The periodic table can help to determine the number of valence electrons in each atom and the central atom is the atom with the greatest number of unpaired valence electrons. **15a:** Valence structures are based on electron dot diagrams. **15b:** The ball-and-stick model is a clearer representation than the space-filling model. **15c:** The representation that one chooses for a molecule depends on its desired purpose. **16a:** The periodic table can help to determine the number of valence electrons in each atom. **16b:** Lone pairs of electrons are those that are not involved in bonding. **16c:** The valence electrons of oxygen can be determined with the help of the periodic table. **17a:** The type of covalent bond depends on the number of electrons shared in the bond. **17b:** The type of covalent bond depends on the number of electrons shared in the bond. **17c:** Double covalent bonds are stronger than single covalent bonds. **18a:** Covalent bonds are formed as a result of atoms wanting to fill their valence shell (the octet rule). **18b:** Covalent bonds are formed as a result of atoms wanting to fill their valence shell (the octet rule). **18c:** Covalent bonds are formed as a result of atoms wanting to fill their valence shell (the octet rule). **18d:** Carbon compounds are diverse due to the carbon atom's ability to form numerous covalent bonds. **19a:** Covalent bonds result from the sharing of electrons. **19b:** Ionic and covalent bonding depends on the type of elements present and their differences in electronegativity. **19c:** Bonding properties can be determined experimentally. **20a:** The independent variable is deliberately manipulated (or changed) by the experimenter. **20b:** The aim is the purpose/objective of an experiment. **20c:** Qualitative data relates to observations while quantitative data relates to numbers. **20d:** Covalently bonded compounds generally have low melting points compared to other types of intramolecular bonding like ionic or metallic bonds.

5B PROPERTIES OF COVALENT COMPOUNDS

The shape and types of covalent bonds present in covalent substances have important effects on a compound's polarity.

5A Covalent bonding | 5B Properties of covalent compounds | 5C Intermolecular bonding

Study design dot point

- shapes of molecules and an explanation of their polar or non-polar character with reference to the electronegativities of their atoms and electron-pair repulsion theory

Key knowledge units

Shapes of molecules	1.2.2.1
Properties determining polarity	1.2.2.2

Key terms and definitions

Valence shell electron pair repulsion theory (VSEPR) model used to predict the different shapes of molecules based on the extent to which electron pairs repel each other

Tetrahedron shape with four triangular faces

Electronegativity how strongly an atom attracts electrons towards itself

Polarity property of having a partial positive and negative charge

Non-polar covalent bond covalent bond where the electrons are shared equally

Polar molecule molecule with a partially positively charged end and a partially negatively charged end

Polar covalent bond covalent bond between atoms with an unequal distribution of electrons in the bond

Permanent dipole moment covalent molecules with a permanent positive and negative charge due to the difference in electronegativity between the atoms

Lesson links

This lesson builds on:

- 5A Covalent bonding

There are different types of covalent bonds which will be explored in this lesson.

Shapes of molecules 1.2.2.1

OVERVIEW

The shapes of molecules are largely determined by the repulsion of electron pairs.

THEORY DETAILS

Valence shell electron pair repulsion theory (VSEPR) provides a way of predicting what the shapes of molecules will be according to the repulsion of electron pairs. Since like charges repel each other, negatively charged pairs of electrons are most stable when they are as far apart as possible. Think about trying to force the negatively charged poles of two magnets together. It is very difficult as the poles repel each other. For most atoms, since covalent compounds are most stable when they have eight electrons in their valence shell according to the octet rule, there are four pairs of electrons in the valence shell.
For example, methane is a hydrocarbon with the molecular formula CH_4. There are four pairs of bonding electrons around the central carbon atom (four electrons originating from the carbon atom and four electrons originating from the four hydrogen atoms). Therefore, these four pairs would like to exist as far apart as they possibly can be. The way that these four pairs of bonding electrons can be positioned as far apart as possible is through a tetrahedral arrangement as shown in figure 1. A **tetrahedron** is the term used in geometry to describe a shape with four triangular faces.

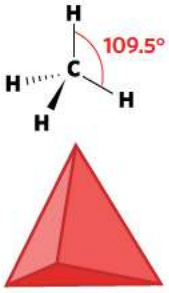


Figure 1 Tetrahedral arrangement in methane (the thick wedge represents a bond that points out of the page, whereas the dashed wedge represents a bond that points into the page).

However, not all molecules have their valence electrons involved in bonding. Some contain lone pairs of electrons which do not take part in bonding. Ammonia has the molecular formula of NH_3 and therefore contains nitrogen and hydrogen atoms.

Nitrogen has five valence electrons and so when bonded to three hydrogen atoms has a complete valence shell with eight electrons. However, one pair of electrons is not involved in bonding. In this case, the four pairs of electrons take a tetrahedral arrangement around the central nitrogen atom as shown in figure 2. Despite this, the shape of the molecule is categorised as trigonal pyramidal because the lone pair of electrons is not counted in the shape of the molecule.

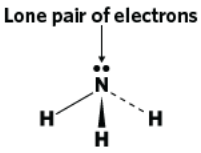


Figure 2 Trigonal pyramidal shape of ammonia.

When there are two lone pairs of electrons present, the shape is again different. The Lewis diagram for water is shown in figure 3.



Figure 3 Lewis diagram of water.



Lone electron pairs are not included in the overall shape of a molecule, only bonding electrons are.

Oxygen has six valence electrons and therefore, when bonded with two hydrogen atoms in order to obtain a stable octet, is left with two lone pairs of electrons. As with the examples before, the pairs of electrons adopt a tetrahedral arrangement around the central oxygen atom resulting in an overall bent or V-shape of the water molecule as shown in figure 4.

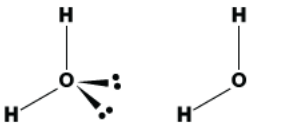


Figure 4 Bent shape of a water.

Three lone pairs of electrons results in a different shape of the overall molecule. The electron dot formula for hydrogen chloride is shown in figure 5. The chlorine atom has seven valence electrons which means that it needs to form a bond with a hydrogen atom to achieve a full valence shell. When chlorine forms a covalent bond with hydrogen, there is one pair of bonding electrons and three lone pairs of electrons (non-bonding electrons). Similar to water, the electron pairs adopt a tetrahedral arrangement about the chlorine atom but the overall shape of the molecule is linear as shown in figure 6.



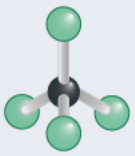
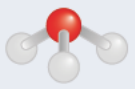
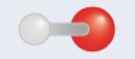
Figure 5 Electron dot diagram of hydrogen chloride.



Figure 6 Linear shape of hydrogen chloride.

The different shapes covalent molecules adopt are summarised in table 1.

Table 1 Different shapes of molecules according to VSEPR.

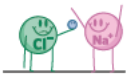
Molecule	Electron dot diagram	Number of lone electron pairs around central atom	Number of single bonds	Ball-and-stick model	Shape
CCl_4		0	4		Tetrahedral
H_3O^+		1	3		Trigonal pyramidal
H_2O		2	2		Bent or V-shaped
HBr		3	1		Linear

So far, we have only considered the shape of covalent molecules that form single bonds as outlined in table 1. When determining the shape of molecules that form double or triple bonds, that is, molecules which contain atoms that share two or three electron pairs respectively, there is an additional rule to consider. This rule states that both double and triple bonds are treated as a single covalent bond and count as one pair of electrons. This gives rise to a range of additional covalent molecular shapes which we will examine more closely.

Methanal (CH_2O) is an organic molecule which consists of a carbon-oxygen double bond and two carbon-hydrogen single bonds. The electron dot diagram of methanal is shown in figure 7.



Figure 7 Electron dot diagram of methanal.



In this case, the carbon-oxygen double bond counts as a single pair of electrons when determining the shape of the molecule. Therefore, if a central atom has three pairs of electrons surrounding it, the electron pairs will adopt positions 120 degrees apart from each other as this is the furthest apart that they can get. This means the shape of methanal is trigonal planar as shown in figure 8.

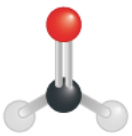


Figure 8 Trigonal planar shape of methanal.

A similar explanation is used when the central atom has two double bonds. Carbon dioxide (CO_2) has two carbon-oxygen double bonds and in this case, since double bonds are treated as single pairs of electrons, the furthest apart that the 'two' pairs of electrons can be is opposite each other, separated by an angle of 180 degrees. Therefore, the molecule will adopt a linear shape as shown in figure 9.



Figure 9 Linear shape of carbon dioxide.

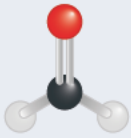
Other covalent molecules such as hydrogen cyanide (HCN) contain a central atom with one triple bond and one single bond. In the case of hydrogen cyanide, the triple bond is treated as one pair of electrons and therefore, since there are 'two' pairs of electrons, each electron pair adopts a position at the opposite end of the central atom, 180 degrees apart, just like in carbon dioxide. This results in a linear shape as shown in figure 10.



Figure 10 Linear shape of hydrogen cyanide.

These different shapes can be summarised in table 2.

Table 2 Shapes of molecules with double and triple bonds according to VSEPR.

Example	Is there a double/ triple bond?	Number of bonds	Ball and stick model	Shape
Methanal	Yes	Single, single, double = 3		Trigonal planar
Carbon dioxide	Yes	Double, double = 2		Linear
Hydrogen cyanide	Yes	Single, triple = 2		Linear

For the sake of this study design, we do not need to know the angle of each of the bonds, just that they are of different shapes.

Another reason why different molecular shapes occur is that not all molecules obey the octet rule. This means that some atoms don't require eight electrons in their valence shell to be stable. Boron is just one example. The shapes of these molecules are not required for the purposes of VCE but can be explained using VSEPR.

Worked example 1

According to VSEPR, what is the shape of NI_3?

What information is presented in the question?

The shape of NI_3 needs to be determined using valence shell electron pair repulsion theory.

What is the question asking us to do?

Describe the shape of NI_3.

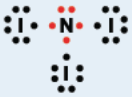
What strategy(ies) do we need in order to answer the question?

1. Determine the electron configuration of N.
2. Determine the electron configuration of I.
3. Determine the types of covalent bonds present and arrange the valence electrons, according to the octet rule.
4. Describe the overall shape of the molecule.

Answer

Nitrogen is in Group 15 of the periodic table and therefore, there are five valence electrons.

Iodine is in Group 17 of the periodic table and therefore, there are seven valence electrons.



Since each atom wants to fill their valence shell with eight electrons according to the octet rule, nitrogen will form a single covalent bond with each of the three iodine atoms. This will enable all atoms to satisfy the octet rule.



Since there are three pairs of bonding electrons and one lone pair of electrons, the overall shape of this molecule is trigonal pyramidal.



Properties determining polarity 1.2.2.2

OVERVIEW

Both the shape of a molecule and the atoms involved determine the polarity of a covalently bonded molecule.

THEORY DETAILS

As we learned in lesson 2C, different atoms have different electronegativities. **Electronegativity** is a measure of how strongly an atom can attract electrons to itself. When two atoms with greatly differing electronegativities form a covalent bond with one another, the atom with the greater electronegativity attracts electrons more strongly towards itself as shown in figure 11.

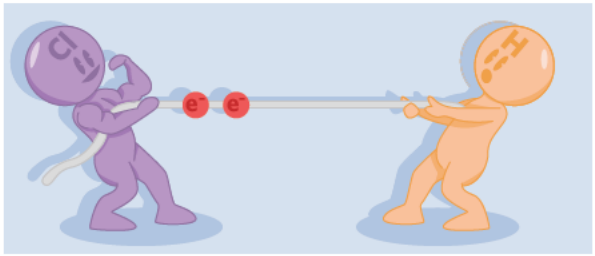


Figure 11 Illustration of chlorine pulling electrons towards itself more strongly than hydrogen.

In the case of hydrogen chloride, chlorine pulls electrons towards itself more strongly than hydrogen due to possessing greater electronegativity. This results in the chlorine atom having a greater electron density around it than the hydrogen atom, which causes chlorine to possess a partial negative charge and hydrogen to possess a partial positive charge. Figure 12 is a common visualisation of this idea with the blue area representing the electron rich region and the red region representing the electron deficient region Notice that the electron cloud associated with chlorine also takes into account the lone pair valence electrons present in the atom. The lowercase Greek letter delta (δ) is used to represent the partial nature of the charge, followed by the '+' or '–' sign to indicate whether the charge is positive or negative. This notation is shown in figure 13.

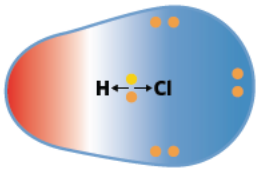


Figure 12 Electron distribution in a polar covalent bond.

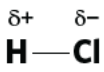


Figure 13 Notation to represent the partial charges associated with HCl.

Polarity in the context of chemistry means that there is a partial positive and partial negative charge present. A polar covalent bond is formed when two atoms with a relatively large difference in electronegativities are covalently bonded together whereas a **non-polar covalent bond** is formed when atoms of the same or similar electronegativity are covalently bonded together. For example, due to the presence of a partial positively and partial negatively charged end, HCl is regarded as a **polar molecule** as well as having a **polar covalent bond**. As there are two 'poles' present in the molecule, we can say that HCl has a **permanent dipole moment**. On the other hand, non-polar covalent bonds usually (but not always) involve two atoms of the same element being covalently bonded together like Cl_2, shown in figure 14.

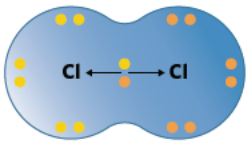
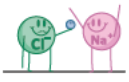


Figure 14 Electron distribution in non-polar covalent bond.



If the difference in electronegativity between each bonded atom is really small, the bond is also regarded as non-polar. This can be seen when a carbon atom is bonded to a hydrogen atom, as shown in figure 15. Table 3 shows how the difference in electronegativity between two atoms can be used to predict the type of bond those atoms form.

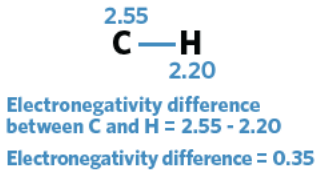


Figure 15 The non-polar covalent bond between carbon and hydrogen.

Table 3 Relationship between electronegativity and the type of bond formed.

Difference in electronegativity	Type of bond
Less than 0.4	Non-polar covalent bond
Between 0.4 and 2.0	Polar covalent bond
More than 2.0	Ionic bond

Since the difference in electronegativity is 0.35 in the C-H covalent bond which is negligible, the bond between a carbon and hydrogen atom is regarded as non-polar. The greater the difference in electronegativity, the more polar the bond. It can also be noticed that any bond which involves atoms with a difference in electronegativity of more than 2.0 results in the formation of a different type of bond, the ionic bond. The nature of ionic bonds is completely different to covalent bonds.

Worked example 2

Using your knowledge of trends in the periodic table, determine whether the C-Br bond is more or less polar than the C-I bond.

What information is presented in the question?

The covalent bonds involved are C-Br and C-I.

What is the question asking us to do?

Compare the polarity of the two covalent bonds.

What strategy(ies) do we need in order to answer the question?

1. Determine which atom, Br or I, is more electronegative than the other.
2. Determine which bond will have the greatest difference in electronegativity.
3. Justify which bond is more polar.

Answer

Bromine is in row 4 of the periodic table whereas iodine is in row 5. Since electronegativity, which is a measure of how strongly an atom attracts electrons towards itself, decreases as one goes down a group on the periodic table, iodine is less electronegative than bromine. Therefore, the difference in electronegativity will be greater in the C-Br bond than it is in the C-I bond. Therefore, the C-Br bond is more polar than the C-I bond.

If a molecule has no polar covalent bonds, it must be non-polar overall. However, the presence of a polar covalent bond does not necessarily mean that the molecule is polar. This is due to the effect that molecular shape has on polarity. Therefore, the second requirement that needs to be taken into account when determining whether a molecule is polar is its VSEPR configuration, an idea we discussed earlier in this lesson. In general, if a molecule is perfectly symmetrical, the partial positive and negative charges created by the difference in electronegativity between atoms cancels out for each bond. For example, carbon tetrachloride has four carbon-chlorine covalent bonds, each of which is polar. However, since the molecule is in a tetrahedral arrangement and symmetrical in three-dimensional space, the partial positive and negative charges created by the differences in electronegativity mean that while each of the covalent bonds is polar, the overall molecule does not display polarity. In other words, there is no overall dipole moment that exists in the whole molecule, as shown in figure 16. Carbon tetrachloride is therefore a non-polar molecule. However, if we change just one of the chlorine atoms into a hydrogen atom, the molecule is no longer symmetrical. Examining the molecules shown in figure 17, we can see that the substitution of a single chlorine for a hydrogen atom results in a partially negative charge in the area associated with the cluster of chlorine atoms and a partially positive charge on the end of the molecule associated with the hydrogen atom. This means that the overall molecule is polar.

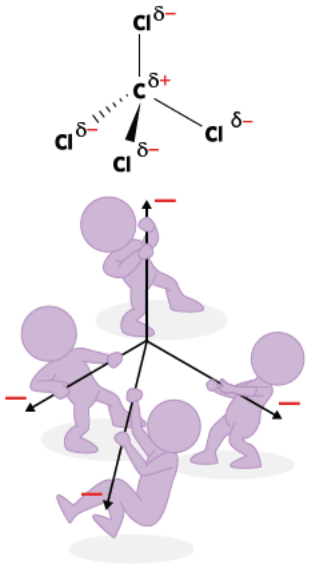
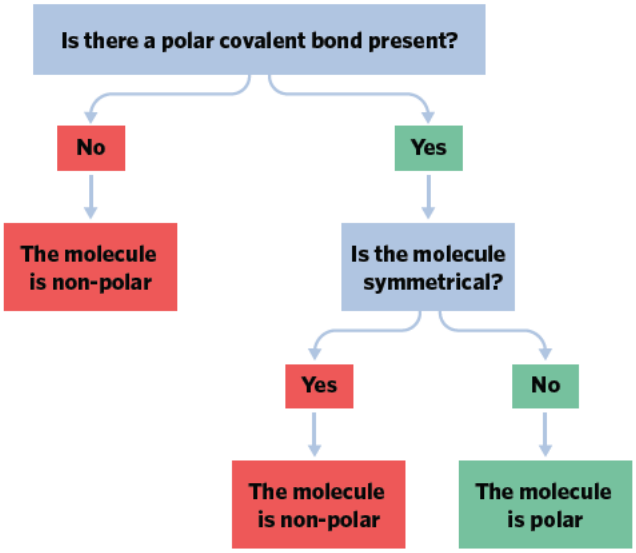


Figure 16 Carbon tetrachloride with the partial positive and negative charges marked.

Here is a flowchart which is helpful in determining the polarity of a molecule:



Whether a molecule is polar or not will have important consequences on the type of intermolecular bonding between molecules that will be explored in the next lesson.

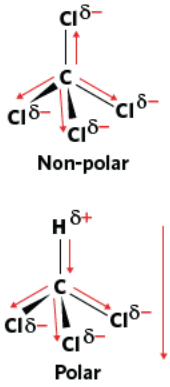


Figure 17 Comparison of a polar and non-polar molecule.



A molecule that contains a polar covalent bond is not necessarily polar.

Worked example 3

HBr and CBr_4 are both compounds which have a highly electronegative bromine atom. One molecule is polar and the other is non-polar. Determine which is polar and which is non-polar and justify your answer.

What information is presented in the question?

The molecules in question are HBr and CBr_4.

What is the question asking us to do?

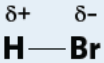
Identify the polar and the non-polar molecule.

What strategy(ies) do we need in order to answer the question?

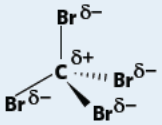
1. Determine whether there are polar covalent bonds present in either molecule.
2. Determine whether the molecules are symmetrical overall.
3. Establish which molecule is polar and which is non-polar.

Answer

Bromine is a highly electronegative element which means that it forms a polar covalent bond with either hydrogen or carbon. HBr has a linear shape which means overall it is a polar molecule.

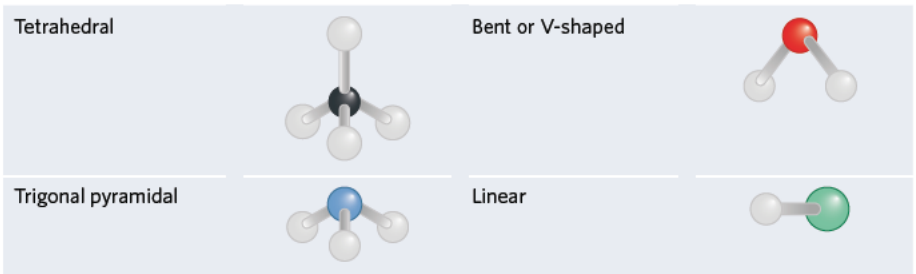


CBr_4 has a tetrahedral shape and even though it consists of four polar covalent bonds, it is non-polar overall due to its symmetry.



Theory summary

- Valence shell electron pair repulsion theory (VSEPR) theory determines the overall shapes of molecules.



- Electron pairs that are involved in bonding contribute to the overall shape of covalent molecules whereas lone electron pairs influence the shape but are not counted in the overall shape.
- Differences in electronegativity determine the polarity of covalent bonds with greater differences in electronegativity leading to more polar covalent bonds.
- The polarity of a molecule depends on both the presence of polar covalent bonds and the molecule's overall shape and symmetry according to VSEPR theory.

5B QUESTIONS

Theory review questions

Question 1

The shape pictured is



A tetrahedral.

B trigonal planar.

Question 2

The shape of PF_3 (pictured) is



A tetrahedral.

B trigonal pyramidal.

Question 3

When it comes to determining the shape of a molecule, a double bond is the same as two single bonds.

A True

B False

Question 4

Which of the following statements is correct about polar covalent bonds?

A Polar covalent bonds cannot occur between atoms of the same species.

B Polar covalent bonds occur whenever highly electronegative atoms are present and bonded to themselves.

Question 5

Non-polar covalent bonds always occur

A between atoms of the same species.

B between atoms of different species.

Question 6

Complete the following sentence

_____ polar molecules have polar covalent bonds and _____ molecules with polar covalent bonds are polar.

A All, not all

B Not all, all

Question 7

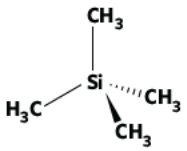
The overall polarity of a molecule depends on

A only its shape.

B the polarity of the bonds present and the shape.

Question 8

Tetramethylsilane (pictured) is commonly used in analytical techniques and has a tetrahedral shape. Is the molecule symmetrical?



A Yes
B No

Question 9

Tom, a chemistry student compared the electronegativity of carbon (2.55) to the electronegativity of silicon (1.90). He concluded that tetramethylsilane was a polar molecule. Is this correct?

A Yes
B No

Question 10

Both carbon dioxide and water have two atoms bonded to a central atom. Which of the molecules is considered non-polar?

A Carbon dioxide
B Water

Deconstructed questions

The following statement is to be used for questions 11–13.

The polarity of a molecule is an important concept in biology because it can determine whether a molecule is able to enter a cell. A student knows that for a particular cell membrane and mechanism of transport, only non-polar molecules are able to enter the cell. Consequently, the student wants to determine whether ammonia ions (NH_4^+) and/or water (H_2O) can enter the cell.

Question 11 (1 MARK)

Polar covalent bonds are present in

A only NH_4^+.
B only H_2O.
C both NH_4^+ and H_2O.
D neither NH_4^+ and H_2O.

Question 12 (1 MARK)

The shapes of the NH_4^+ molecule and H_2O respectively are:

A Tetrahedral, tetrahedral
B Tetrahedral, bent
C Tetrahedral, linear
D Trigonal pyramidal, linear

Question 13 (4 MARKS)

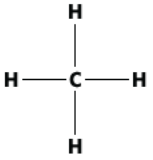
Using the information, determine which molecule will be able to enter the cell and which one will not not be able to enter the cell.

Exam-style questions

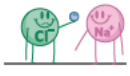
Within lesson

Question 14 (1 MARK)

Why might the following image pose a problem for a student's understanding of VSEPR?



A It is a three-dimensional rather than two-dimensional drawing.
B It appears that the bond angle is 90 degrees rather than 109.5 degrees.
C The drawing is an incorrect structural formula.
D It is too complicated.



Question 15 (1 MARK)

Which of the following options best represents the three-dimensional structure of a molecule according to VSEPR?

A Electron dot diagram
B Valence structure
C Structural formula
D Ball and stick model

Question 16 (3 MARKS)

The figures shown represent the electron clouds surrounding atoms of two different diatomic molecules.

Which of the two molecules would be considered as polar? Explain.



Question 17 (5 MARKS)

Carbon monoxide (CO) and carbon dioxide (CO_2) are both composed of the same types of atoms. However, they have different polarities. Describe the difference in polarity between both molecules with reference to VSEPR and diagrams.

Multiple lessons

Question 18 (8 MARKS)

Phosphine (PH_3) was first obtained in 1783 by a student of the famous French chemist Antoine Lavoisier. Its Material Safety Data Sheet (MSDS) is shown.



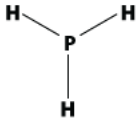
Hazards:

- Extremely flammable gas.
- Contains gas under pressure; may explode if heated.
- Fatal if inhaled.
- Causes severe skin burns and eye damage.

a Draw an electron dot diagram and valence structure of phosphine and label which is which. (2 MARKS)

Antony decided to draw the phosphine molecule to help him investigate some of its properties.

He therefore concluded that according to VSEPR, phosphine was a trigonal planar molecule and since it is symmetrical, it is non-polar.



b Is Antony's conclusion correct? Justify your answer. (4 MARKS)

c According to the MSDS, suggest two safety recommendations when it comes to the storage of phosphine gas. (2 MARKS)

Question 19 (11 MARKS)

Atoms have the ability to bond with other atoms to become molecules. Different molecules can take a different shape depending on the atoms involved.

a Complete the following table. (4 MARKS)

Number of lone pairs on central atom	Number of single bonds	Ball and stick model	Shape
1			
	2		Bent or V-shaped

b A student remembers that he once saw a molecule in a trigonal planar shape with no double covalent bonds present. Give an example of how this might be possible. (2 MARKS)

c Is there only one combination of bonds and lone pairs of electrons that gives rise to a linear shape? Justify your answer. (2 MARKS)

d What causes the different shapes of molecules according to VSEPR? (1 MARK)

e Are lone pairs important in determining the shape of the molecule? Explain why or why not. (2 MARKS)

Key science skills

Question 20 (3 MARKS)

A student, Katherine, wanted to test the different effects that the VSEPR shape of a molecule had on its physical and chemical properties. She decided to use HCl (an example of a linear molecule) and CH_4 (an example of a tetrahedral molecule). HCl was a lot more reactive than CH_4. Therefore, she concluded that linear molecules are more reactive than tetrahedral molecules.

a What is the intended independent variable in this experiment? (1 MARK)

b Are the results of this experiment valid? Explain why or why not. (2 MARKS)

Exam-style question hints

14: The two dimensional drawing does not represent the true nature of the geometry of the molecule. **15:** The only representations which support a three-dimensional view of a molecule according to VSEPR are the ball-and-stick and space-filling models. **16:** Polarity depends on the distribution of electron density. **17:** Polarity is affected by molecular shape. **18a:** Electron dot diagrams depict both the lone pairs of electrons and bonding electrons. **18b:** The shape of a molecule is also influenced by the presence of lone pairs of electrons. **18c:** Safety recommendations depend on the specific hazards posed by a chemical. **19a:** Different numbers of single bonds and lone pairs lead to different shapes. **19b:** Some atoms don't obey the octet rule. **19c:** Lone pairs of electrons are not counted when determining the final shape of a molecule. **19d:** Like charges repel each other. **19e:** The number of pairs of electrons (both bonding and lone pairs) helps to determine the shape of the molecule. **20a:** The independent variable is what is being changed in each experiment. **20b:** Validity is a measure of whether the experiment can prove the aim.

5C INTERMOLECULAR BONDING

In the previous lessons, we learned about the bonds that hold atoms together to form molecules. In this lesson, we will explore the bonds that hold molecules together.

5A Covalent bonding | 5B Properties of covalent compounds | 5C Intermolecular bonding

Study design dot points

- the relative strengths of bonds (covalent bonding, dispersion forces, dipole-dipole attraction and hydrogen bonding) and evidence and factors that determine bond strength including explanations for the floating of ice and expansion of water at higher temperatures
- explanation of properties of molecular substances (including low melting point and boiling point, softness, and non-conduction of electricity) with reference to their structure, intramolecular bonding and intermolecular forces

Key knowledge units

Dispersion forces	1.2.4.2
Dipole-dipole attraction	1.2.4.3
Hydrogen bonding	1.2.4.4
Properties of molecular substances	1.2.3.1

Key terms and definitions

Orbital region around the nucleus of an atom where electrons are likely to be found

Instantaneous dipole moment partial positive and negative charge that suddenly appears in an atom/molecule due to the random movement of electrons

Dispersion force electrostatic force of attraction between molecules resulting from instantaneous dipole moments

Intermolecular forces forces that attract molecules to each other

Dipole-dipole force electrostatic force of attraction between a partial positive charge on one molecule and a partial negative charge on another molecule

Melting point temperature at which a substance changes its state from solid to liquid

Boiling point temperature at which a liquid transitions into a gas

Kinetic energy energy of an object due to its motion

Molecular substance substance with two or more atoms joined by covalent bonds

Hardness resistance to deformation when subjected to pressure

Lesson links

This lesson builds on:

- 2A The atom

An understanding of quantum mechanics is important in explaining some intermolecular forces.

- 5B Properties of covalent compounds

The polarity of molecules determines the type of intermolecular forces present.

Dispersion forces 1.2.4.2

OVERVIEW

All molecules are attracted to each other by weak forces called dispersion forces.

THEORY DETAILS

As we learned in lesson 2A, electrons occupy regions which are called orbitals. **Orbitals** represent the area where an electron is likely to be found and do not represent the exact location that an electron will be. This concept is illustrated in figure 1. The grey dots represent the probability of finding an electron in a specific region, with the majority of these dots centred in the middle. This means that an electron in the 1s subshell is most likely to be found closer to the centre of an atom. However, at a random point in time, the electron(s) in the 1s subshell might be located on either the far right or far left side of the central region, as shown in figure 2.



Figure 1 Probability distribution of the 1s orbital of an atom.

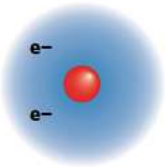


Figure 2 Diagram of an atom with the electrons on the far left of the orbital.

When electrons spontaneously move to the far left or right of an atom, an **instantaneous dipole moment** (also known as a spontaneous dipole moment) is created. Since electrons are negatively charged, and are located on the left-hand side, this side has a partial negative charge and the right-hand side, without any electrons, has a partial positive charge. This polarity is depicted in figure 3.

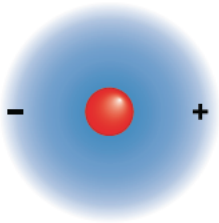


Figure 3 Diagram of an atom with an instantaneous dipole-dipole moment.

Since positive and negative charges attract, sometimes another atom with a similar instantaneous dipole moment can be nearby, resulting in an intermolecular force of attraction known as a **dispersion force**. This scenario is shown in figure 4.

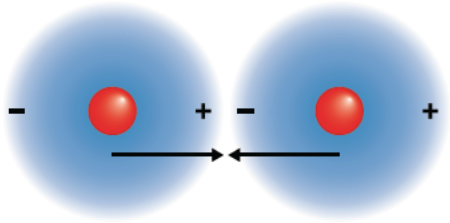


Figure 4 Instantaneous dipole-dipole moment between two atoms with arrows representing the force of attraction.

It should be highlighted that dispersion forces are the weakest type of **intermolecular force**. This is because the electrons, which are randomly distributed in the orbitals, have to line up exactly in order to create an instantaneous dipole-dipole moment. Since electrons are continuously in motion, the instantaneous dipole-dipole moment formed is momentary, and the dispersion forces will be disrupted once these electrons shift to a different part of the orbital.

The most important factor for determining the strength of dispersion forces is the number of electrons. While in the examples given above, we have looked exclusively at atoms, the exact same reasoning applies to molecules which consist of atoms chemically bonded together. The more electrons a molecule has, the greater the possibility that an instantaneous dipole-dipole moment will be created. This is because for larger atoms and molecules, there are more electrons that can form instantaneous dipoles. With more instantaneous dipoles, the probability of an instantaneous dipole-dipole moment between molecules increases which contributes to an increase in the strength of dispersion forces. As a result, a molecule with more electrons will have stronger dispersion forces than a molecule with fewer electrons. For example, methane (CH_4) has weaker dispersion forces than octane (C_8H_{18}) as it has fewer electrons (10 electrons compared to 66 electrons). Let's explore the concept further with an example.

Worked example 1

Methane and ethane are both hydrocarbons. Methane has the molecular formula CH_4 and ethane has the molecular formula C_2H_6. Which molecule would have the strongest dispersion forces and why?

What information is presented in the question?

The molecular formulas of methane and ethane.

What is the question asking us to do?

Compare the strength of dispersion forces of methane and ethane.

What strategy(ies) do we need in order to answer the question?

1. Use the periodic table to determine the number of electrons in each molecule.
2. Use the information from step 1 to determine which molecule has the strongest dispersion forces.
3. Answer the question with a justifying statement.

Answer

According to the periodic table, carbon has six electrons (atomic number 6) and hydrogen has one electron (atomic number 1).

Therefore, methane (CH_4) has $1 \times 6 + 4 \times 1 = 10e^-$ and ethane (C_2H_6) has $2 \times 6 + 6 \times 1 = 18e^-$.

Since ethane has more electrons than methane and the strength of dispersion forces is proportional to the number of electrons, ethane has stronger dispersion forces than methane.

Dipole-dipole attraction 1.2.4.3

OVERVIEW

Dipole-dipole attractions are another type of intermolecular force that are formed from permanent dipole-dipole moments rather than temporary dipole-dipole moments.

THEORY DETAILS

As we learnt in lesson 5B, molecules can be classified as either polar or non-polar. Polar molecules have a permanent dipole moment because there is a permanent partial positive and negative end of the molecule. Instantaneous dipole moments, however, arise due to the random movement of electrons within molecules and create dispersion forces. The key difference between permanent **dipole-dipole forces** and dispersion forces is that dispersion forces result from an instantaneous dipole moment, whereas permanent dipole-dipole forces always have a dipole moment, as shown in figure 5.



Figure 5 Permanent dipole-dipole forces between HCl molecules.

The highly electronegative chlorine atom in the HCl molecule readily attracts electrons which gives chlorine a partial negative charge. Since electrons have been drawn away from the hydrogen atom, it gains a partial positive charge. These partial negative and positive charges are denoted by δ− and δ+ respectively. As shown in figure 5, there will be an attraction between the δ− and δ+ dipoles of different HCl molecules.

The partial positive and negative charges on molecules influences the strength of the attractive forces between molecules because permanent dipole-dipole bonds are stronger than dispersion forces. The strength of permanent dipole-dipole forces also depends on the difference in electronegativity of the atoms involved in the bond.

Worked example 2

HBr and CBr_4 both have a highly electronegative bromine atom. With respect to their intermolecular forces of attraction, which molecule would have the strongest permanent dipole-dipole bonds?

What information is presented in the question?

The molecules in question are HBr and CBr_4.

What is the question asking us to do?

Compare the strength of each molecule's dipole-dipole bonds.

What strategy(ies) do we need in order to answer the question?

1. Compare the electronegativities of the atoms covalently bonded together to determine whether there are polar covalent bonds present in either molecule.
2. Determine whether the molecules are polar overall using symmetry.
3. Answer the question with a justifying statement.

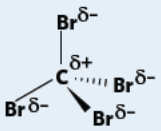
Answer

Bromine is a highly electronegative element which forms polar covalent bonds with hydrogen or carbon.

HBr has a linear shape which means overall it is a polar molecule.



CBr_4 has a tetrahedral shape and even though it consists of four polar covalent bonds, it is non-polar overall due to symmetry.



Therefore, since HBr is polar, it can form permanent dipole-dipole forces, whereas CBr_4 is non-polar, meaning it can only form dispersion forces (instantaneous dipole-dipole forces). Hence HBr has the strongest permanent dipole-dipole interactions.

Hydrogen bonding 1.2.4.4

OVERVIEW

Hydrogen bonding is a type of permanent dipole-dipole bonding and only occurs in molecules when hydrogen is bonded to one of fluorine, oxygen or nitrogen.

THEORY DETAILS

Hydrogen bonds are a type of permanent dipole-dipole bond that are considerably stronger than instantaneous dipole-dipole bonds. This is because the polarity of the bond that is involved is much greater due to hydrogen, a weakly electronegative atom, being bonded to one of the strongly electronegative fluorine, oxygen or nitrogen atoms as shown in figure 6. This means that there is a stronger electrostatic force of attraction between the partially positive hydrogen atom and the partially negative oxygen, fluorine or nitrogen atom.

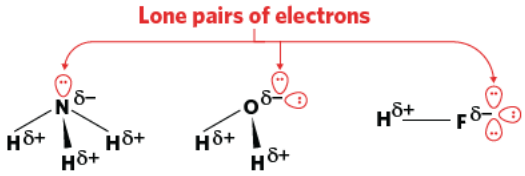


Figure 6 Diagram illustrating the permanent dipole moments of the N-H, O-H and F-H bond.

The pairs of bonded atoms in figure 6 create a very strong dipole moment which increases the strength of the permanent dipole-dipole forces, enough for chemists to classify the dipole forces as hydrogen bonds. Figure 7 illustrates that for a hydrogen bond to form, the hydrogen atom has to be positioned between two highly electronegative atoms (fluorine, oxygen or nitrogen) where one of those electronegative atoms is from a different molecule and has its lone pair of electrons free.

Table 1 summarises the relative strength of dispersion forces, permanent dipole-dipole bonds and hydrogen bonds.

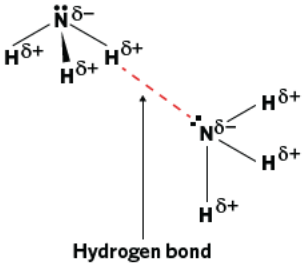


Figure 7 Hydrogen bonding between ammonia molecules.

Table 1 Relative strength of dispersion forces, permanent dipole-dipole bonds and hydrogen bonds.

Intermolecular force	Relative strength
Dispersion forces	1
Permanent dipole-dipole forces	10
Hydrogen bonds	50

The strength of hydrogen bonds gives water its special properties which include ice's ability to float and the expansion of water as it is heated. This is because intermolecular bonding is able to explain certain physical properties. For example, a molecule which has very strong intermolecular forces will have a higher **melting point** and **boiling point** than a molecule with weaker intermolecular forces because it requires more energy to break apart the molecules. Let's apply this concept to water. Figure 8 illustrates the hydrogen bonds that form between water molecules.

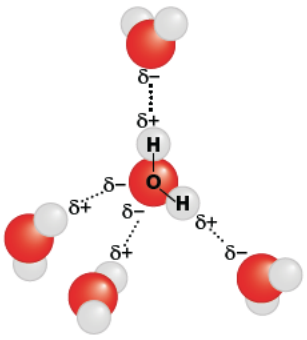


Figure 8 Hydrogen bonding of water molecules.

Due to the strength of hydrogen bonding, water has a very high boiling point relative to the size of the molecule. Resulting from water's special structure, water can form four hydrogen bonds with other water molecules as shown in figure 8. When the temperature of water reaches its freezing point at 0°C, the water molecules start to arrange themselves into a regular lattice, with each water molecule bonded to four others, which means they are spaced further apart from each other as shown in figure 9.

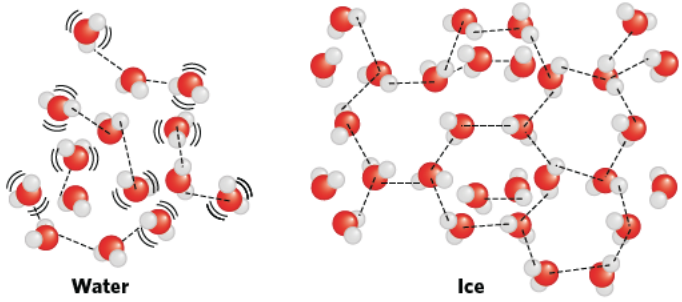


Figure 9 Arrangement of water molecules when in the liquid state (left) compared to ice, the solid state (right).

The arrangement of water molecules into a regular lattice at freezing temperature means that ice has a lower mass per volume (also known as density) than liquid water. This explains why ice is able to float on water despite both water and ice consisting of the same molecule. The relationship between the temperature and volume of liquid water is shown in figure 10.

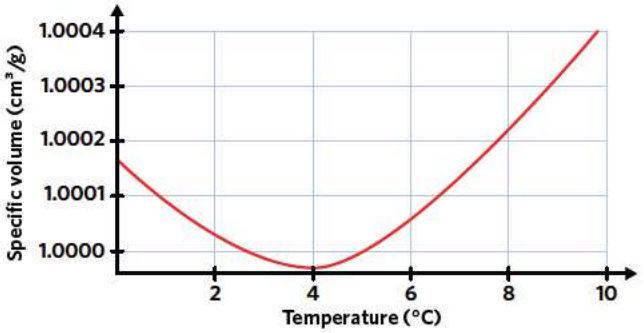


Figure 10 Graph illustrating the relationship between the temperature and volume of liquid water.

Past 4°C, water expands when heated. This is because since water molecules have greater **kinetic energy** at higher temperatures, they start to move and vibrate more resulting in a less dense arrangement as they become more dispersed and the hydrogen bonding between the water molecules is disrupted.

Properties of molecular substances 1.2.3.1

OVERVIEW

There are a number of different factors such as the identity of atoms present which affect the properties of covalent molecular substances.

THEORY DETAILS

For the most part, the properties of **molecular substances** are determined by their intermolecular forces – the forces that exist between molecules. This is expressed with the prefix inter (between) as opposed to intra (within). Intermolecular forces are responsible for a number of different physical properties including melting point, boiling point and **hardness**.



Image: optimarc/Shutterstock.com

Figure 11 Insect walking on the surface of water.

Evidence of intermolecular forces is all around us. For example, some insects are able to walk on the surface of water, as shown in figure 11. This phenomenon is due to a concept known as surface tension. The strong intermolecular hydrogen bonding between water molecules allows the surface of the water to support the weight of the insect. Without the surface tension and intermolecular bonding, this would not be possible. Hardness is another example of the effect of the strength of intermolecular forces on physical properties. Soft substances like butter and wax are covalent molecular substances where the intermolecular forces are much weaker than those found in ionic compounds or metals.

The strength of intermolecular forces reflects the amount of energy it takes to break the bonds between molecules. At each phase of matter, solid, liquid and gas, molecules have a certain amount of kinetic energy. This kinetic energy allows the molecules in the substance to move around and overcome the forces that are maintaining the substance in a particular state.

For example, molecules that exist in the solid state have low kinetic energy and therefore cannot overcome the intermolecular forces that are holding them together. When there is enough energy to overcome these forces, the substance is able to transition into a liquid, which has weaker intermolecular forces. The temperature at which this occurs is known as the melting point. Finally, with enough kinetic energy, molecules can break all intermolecular forces to exist as a gas. When substances change from a liquid to a gaseous state, this is known as the boiling point which is shown in figure 12.

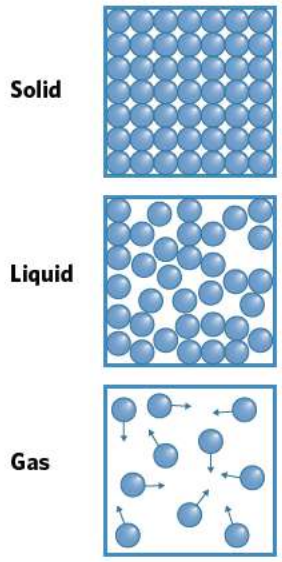


Figure 12 Arrangement of molecules in different states of matter.

As we know, different intermolecular bonds vary in strength. When compared to each other, hydrogen bonds are the strongest intermolecular forces, followed by permanent dipole-dipole interactions and finally dispersion forces.

This means that it takes more energy to break hydrogen bonds than permanent dipole-dipole forces, which in turn require more energy to break apart than dispersion forces. Therefore, the types of intermolecular bonding present between molecules in a substance plays an important role in determining the properties of that particular molecular substance.

The size and structure of molecules also play an important role in determining the strength of intermolecular forces. Take for example a substance made from only non-polar covalent bonds. Since there are no highly electronegative atoms, this substance will be held together by dispersion forces only - the weakest of the three intermolecular forces. However, as the molecule becomes larger, the overall strength of the dispersion forces increases, resulting in a greater melting point. This interaction is shown in figure 13.

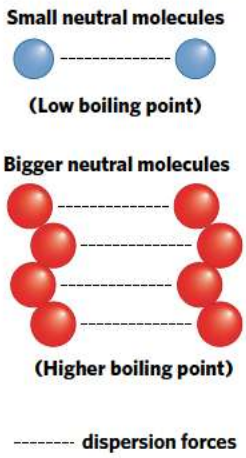


Figure 13 The impact of molecular size on the strength of dispersion forces.

Molecules need access to each other in order to form intermolecular bonds and molecules with a greater surface area form intermolecular bonds more easily than molecules with a small surface area. As a result, molecules with a greater surface area form stronger intermolecular bonds. An example of this involves the toes of geckos. Geckos are able to cling to both polar and nonpolar surfaces because of dispersion forces. Since dispersion forces are very weak, the toes of geckos have millions of little hairs which increase the surface area of the gecko. This creates enough intermolecular forces of attraction for geckos to cling onto walls as shown in figure 14.

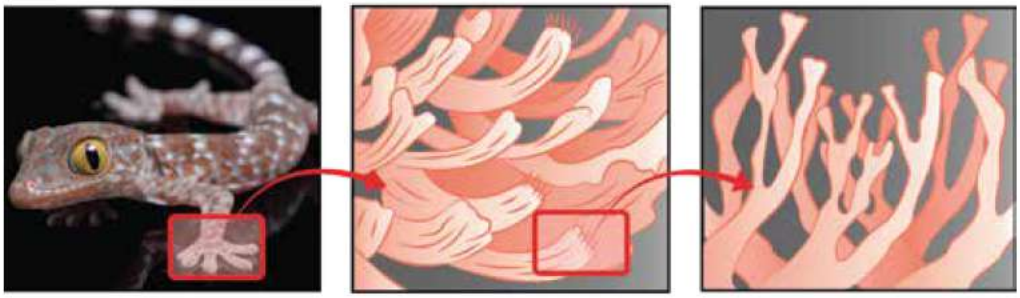


Image: dwi putra stock/Shutterstock.com

Figure 14 Toes of geckos.

Another factor that determines the strength of intermolecular forces is the distance between molecules. Molecules which can pack closely together will have stronger intermolecular forces than molecules which cannot pack as closely. For example, this explains why a heavily branched molecule will have weaker intermolecular forces than a linear molecule as shown in figure 15.

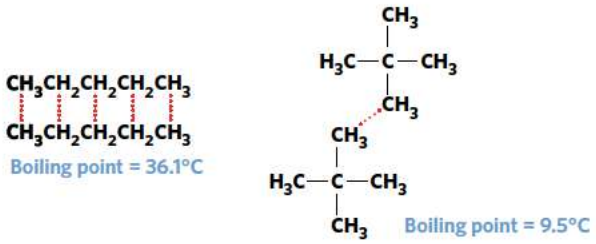
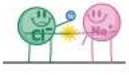


Figure 15 Comparison of boiling points between isomers pentane and dimethylpropane.



The melting and boiling points of a substance are a reflection of the strength of the intermolecular bonds and not the intramolecular bonds.

The strength of intermolecular bonding is approximately 100 times less than that of intramolecular bonding like covalent bonds or ionic bonds. Using water as an example, it is relatively easier to break the intermolecular bonds between water molecules as opposed to the intramolecular bonds in water, which are the covalent bonds between oxygen and hydrogen atoms. To break the covalent (intramolecular) bonds in water, it requires a temperature of approximately 3000°C, whereas 100°C is enough to break the hydrogen (intermolecular) bonds between water molecules.



Worked example 3

Explain why methanol (CH_3OH) has a higher boiling point than methane (CH_4) with reference to intermolecular forces.

What information is presented in the question?

The molecules in question are CH_3OH and CH_4.

What is the question asking us to do?

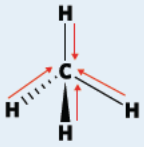
Compare the difference in boiling points, explaining our answer with respect to intermolecular forces.

What strategy(ies) do we need in order to answer the question?

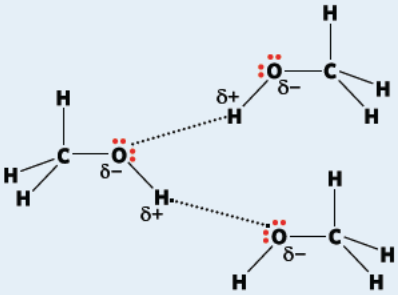
1. Identify the intermolecular forces present in each molecule.
2. Relate the strength of intermolecular forces to the boiling point.
3. Answer the question with a justifying statement.

Answer

Methane is a non-polar molecule as it is symmetrical.



Methanol is a polar molecule because it has a polar covalent bond that occurs between the oxygen and hydrogen in the OH group. This means that the highly electronegative oxygen atom is pulling electrons away from the hydrogen atom. As a result, there is a permanent dipole-dipole moment. Furthermore, the partially positive hydrogen atom on one molecule is able to form a hydrogen bond with a partially negative oxygen atom on a different molecule.



Although both molecules have dispersion forces of roughly equal strength, methanol has hydrogen bonding as well. As a result, methanol will have stronger intermolecular bonding and, therefore, a higher boiling point as more energy is required to break the bonds between the methanol molecules.

Theory summary

- Dispersion forces are caused by instantaneous dipole-dipole moments created by the random movement of electrons.
- The more electrons a molecule has, the stronger the dispersion forces.
- Permanent dipole-dipole forces result from the attraction between partially positive and partially negative charges in polar molecules.
- Permanent dipole-dipole forces are stronger than dispersion forces which only result from instantaneous dipole-dipole moments.
- Hydrogen bonds are a type of permanent dipole-dipole bond and are only present when a partially positive hydrogen atom is bonded to a partially negative fluorine, oxygen or nitrogen atom.
- Ice is less dense than water due to the crystal lattice structure the water molecules adopt at freezing point.
- Water expands when heated as the strong hydrogen bonds begin to be disrupted by the increase in kinetic energy of water molecules.
- Physical properties such as boiling point, melting point and hardness can be determined by the strength of the intermolecular forces present.
- The strength of the intermolecular forces present depends on the type of intermolecular forces and the size and shape of molecules.
- Intermolecular forces of attraction are weaker than intramolecular (e.g. covalent) bonds.

5C QUESTIONS

Theory review questions

Question 1

Intermolecular bonding helps to

A hold atoms together.

B keep molecules together.

Question 2

When determining the strength of intermolecular forces, it can be useful to determine the

A number of electrons.

B number of neutrons.

Question 3

Dispersion forces arise from

A permanent dipole moments caused by polar covalent bonds.

B instantaneous dipole moments caused by randomly moving electrons.

Question 4

Dispersion forces

A exist between all molecules.

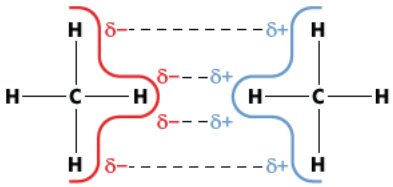
B only exist between small molecules.

Question 5

What does the image shown represent?

A Hydrogen bonds

B Dispersion forces



Question 6

Permanent dipole-dipole forces are _____ than dispersion forces and occur when _____ molecules are present.

A stronger, polar

B weaker, non-polar

Question 7

Given that fluorine is more electronegative than chlorine, would hydrogen fluoride or hydrogen chloride have the strongest intermolecular forces?

A Hydrogen fluoride

B Hydrogen chloride

Question 8

Water has remarkable properties including a high boiling point for its size because

A it has no dispersion forces.

B it has hydrogen bonding.

Question 9

In comparison to breaking intramolecular bonds, the energy required to break intermolecular bonds is

A greater.

B smaller.

Question 10

Molecules with a _____ surface area have _____ interactions with other molecules.

A small, stronger

B large, stronger

Deconstructed questions

The following information is to be used for questions 11–13.

Depending on environmental conditions, molecules can exist in many different forms. For example, water on the planet of *55 Cancri e*, has an average temperature of 1400°C and therefore exists predominantly as vapour. Water, however, exists on earth predominantly in liquid form.

Consider the following statement:

At room temperature, F_2 and Cl_2 are gases, Br_2 is a liquid, and I_2 is a solid.

Question 11 (1 MARK)

The intermolecular forces acting between the different molecules are

A dipole-dipole interactions.

B hydrogen bonds.

C dispersion forces.

D ionic bonds.

Question 12 (1 MARK)

The main difference between these molecules is

A the type of intramolecular bonds holding the molecule together.

B the type of intermolecular bonds holding the molecule together.

C the size of the atoms.

D the strength of the intramolecular bonds holding the molecule together.

Question 13 (3 MARKS)

Explain why the molecules above exist in different physical states at room temperature.

Exam-style questions

Within lesson

Question 14 (9 MARKS)

Trichloromethane, $CHCl_3$, also known as a chloroform, was a common sedative used up until the 1960s.

a Draw the three-dimensional structure of trichloromethane, labelling any partial positive and negative charges, as well as the overall dipole of the molecule. (2 MARKS)

b List all the intermolecular forces that are present in liquid trichloromethane and explain how these forces arise. (4 MARKS)

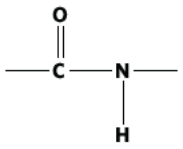
c Chloroform is a dense, yet volatile liquid. With respect to the strength of the intermolecular forces present, justify why it is a liquid at room temperature. (3 MARKS)

Question 15 (6 MARKS)

Proteins are large organic molecules that are important in biology and their function is heavily linked to their structure. One of the important features of proteins is their ability to form hydrogen bonds to stabilise their structure.

a Why might hydrogen bonds as opposed to dispersion forces, be important in stabilising protein structure? (2 MARKS)

The bond that connects the components of a protein together is called the peptide bond and its structure is shown. It consists of a carbon-oxygen double bond and a nitrogen-hydrogen single bond.



Peptide bond

b Draw a possible hydrogen bonding interaction between two different peptide bonds and label the hydrogen bond. (2 MARKS)

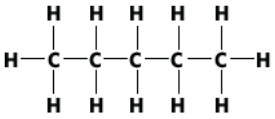
c Suggest one factor that could disrupt the stability of the interactions between two peptide bonds. How do you think this would affect the function of the protein? (2 MARKS)

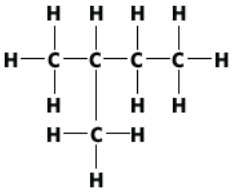
Question 16 (11 MARKS)

A scientist wanted to determine the properties of two different molecules.

a Identify the main intermolecular force present in both molecules. (1 MARK)

b Describe how this intermolecular force arises. (1 MARK)


Molecule A


Molecule B

Unfortunately, the melting point determination results got mixed up and the scientist could no longer determine which sample consisted of which molecule. The results of the analysis are shown:

	Sample 1	Sample 2
Melting point (°C)	−130	−160

c Based on your knowledge of the strength of intermolecular forces, identify which molecule made up each sample. Justify your answer. (5 MARKS)

d Other than melting point, give two other physical properties that are affected by the strength of intermolecular forces and explain this connection. (4 MARKS)

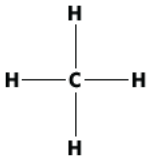
Question 17 (10 MARKS)

Water is one of the most important substances in the world. Not only is it required to maintain bodily functions, it is also involved in many agricultural and material production processes.

a Draw and label the interaction that occurs between four water molecules in the liquid state. (2 MARKS).

b Why is ice able to float on water? Justify your answer with reference to the change in interaction between water molecules. (3 MARKS)

Consider the molecule shown. It has a similar molecular weight to water, however, the melting point of this molecule is −182°C whereas the melting point of water is 0°C.



c Explain this difference in melting points with respect to the intermolecular forces present. (3 MARKS)

d A student predicted that tetradecane ($C_{14}H_{30}$) which only has dispersion forces would also have a lower melting point than water. This, however, is not the case. Explain how this is possible. (2 MARKS)

Question 18 (7 MARKS)

Liquid nitrogen is currently used in the food industry to create ice cream. Liquid nitrogen can be extremely dangerous when not handled correctly and therefore non-scientists working with this substance need to take extra precautions.

Below is some information about liquid nitrogen:

- At atmospheric pressure, liquid nitrogen boils at −195.8°C.
- Nitrogen is relatively inert and inflammable.
- Liquid nitrogen can cause severe frostbite if it comes into contact with living tissue.

a Identify the predominant force of attraction between gaseous nitrogen (N_2) molecules. (1 MARK)

b Describe the strength of the intermolecular bonds between nitrogen molecules. (3 MARKS)

Angela was using liquid nitrogen to make chocolate ice cream. Without looking, she accidentally put her finger into a canister full of liquid nitrogen. However, she was able to remove it from the sample in a split second. When she examined her finger, she noticed that it looked like there was 'smoke' coming off her finger even though she was not injured in any way.

c Explain how it was possible that this occurred and how she was not injured. (3 MARKS)

Multiple lessons

Question 19 (13 MARKS)

Solid sodium chloride (NaCl) is an ionic compound with many different uses. One of the uses of sodium chloride is as a condiment that helps to preserve and flavour food. When placed in a beaker of water, sodium chloride conducts electricity. One of the main reasons why substances can conduct electricity is that they can transfer 'charge' through the movement of charged particles.

a Is ionic bonding in a solid crystal of NaCl an intermolecular or intramolecular force? Explain your answer. (2 MARKS)

b Describe what happens to sodium chloride when it is placed in water. (3 MARKS)

c Explain why solid sodium chloride is unable to conduct electricity unless it is placed in water. (3 MARKS)

d HCl has a melting point of −114°C, whereas NaCl has a melting point of 801°C. Explain this difference with respect to the intermolecular and/or intramolecular forces present. (3 MARKS)

e Similar to NaCl, magnesium chloride ($MgCl_2$) is an ionic compound with many applications. Magnesium chloride is produced through a chemical reaction between solid magnesium (Mg) and chlorine gas (Cl_2).

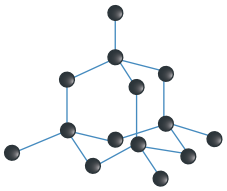
i Write a balanced equation to show the production of magnesium chloride. (1 MARK)

ii If two moles of magnesium react with chlorine gas, how many moles of magnesium chloride will be produced? (1 MARK)

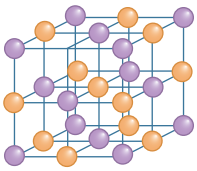
Key science skills

Question 20 (5 MARKS)

Kath and Kim wanted to conduct an experiment to see the differences in boiling points of two different substances containing the two structures shown.



Substance A



Substance B

It is known that substance A is a covalent compound, whereas substance B is an ionic compound. When conducting the experiment, the students used 5 g of substance A and 4 g of substance B.

The results of the experiment are shown.

	Substance A	Substance B
Boiling point (°C)	145	120

a Identify the independent variable. (1 MARK)

b Write an appropriate aim for the experiment. (1 MARK)

c Comment on the validity of the results of the experiment. (3 MARKS)

Exam-style question hints

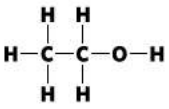
14a: There are four atoms around the central carbon meaning it will adopt a tetrahedral shape. **14b:** Dispersion forces are present in every molecule. **14c:** State of matter and intermolecular forces are interrelated. **15a:** Hydrogen bonds and dispersion forces differ greatly in strength. **15b:** A hydrogen bond occurs between a hydrogen atom bonded to one of F, O, N and the lone pair of electrons on another highly electronegative atom. **15c:** Intermolecular forces can be disrupted with temperature. **16a:** The types of covalent bonds present in a molecule can help to determine the intermolecular forces present. **16b:** Dispersion forces arise from instantaneous dipoles. **16c:** The melting point of a substance depends on the intermolecular bonds present. **16d:** Many of the common physical properties that chemists measure like boiling point and hardness are based on the strength of the intermolecular forces. **17a:** Water is a highly polar compound. **17b:** Water molecules in ice adopt a crystalline structure. **17c:** Melting point is affected by the type(s) of intermolecular bonds present. **17d:** Dispersion forces vary in strength depending on the number of electrons present. **18a:** The intermolecular forces that are present can be determined by the types of intramolecular bonds in the molecule. **18b:** Different intermolecular forces have different strengths. **18c:** A human's body temperature is around 36–37°C. **19a:** One solid crystal of NaCl is one molecule. **19b:** Positive and negative charges attract each other. **19c:** Electricity can only be conducted when there is a movement of charged particles. **19d:** Ionic compounds have very strong intramolecular forces. **19ei:** Chemical equations must be balanced. **19eii:** We can use the mol ratios established in a chemical equation to calculate amounts of different species in a reaction. **20a:** The independent variable is the variable that is manipulated. **20b:** The aim indicates the goal of an experiment. **20c:** The validity of an experiment is affected by the variables changed in the experiment.

EXPERIMENT

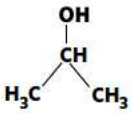
COMPARISON OF THE BOILING POINTS OF MOLECULAR COMPOUNDS

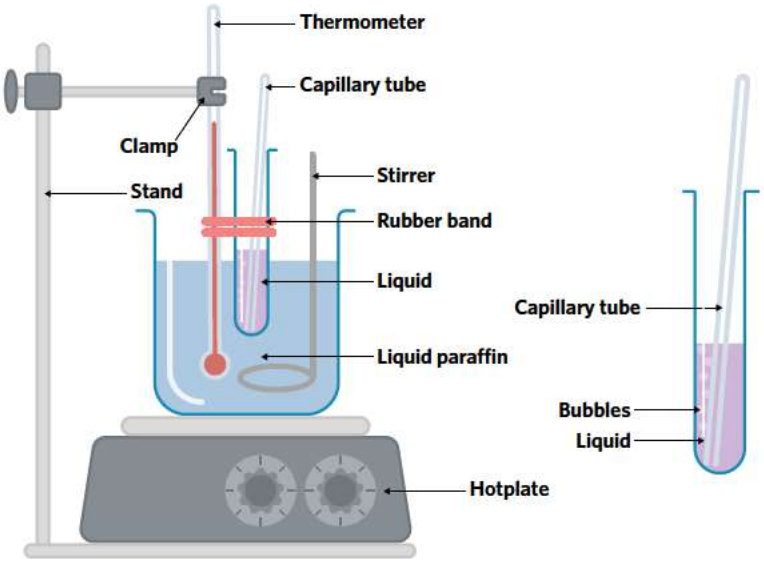
In this chapter, we have investigated covalent bonding and the different intermolecular forces present in a compound. A student decided to conduct an experiment to determine the boiling points of the following two molecules.

Ethanol



Propan-2-ol





Materials

(per group)

- 250 mL glass beaker
- 150 mL liquid paraffin
- 20 mL 0.5 M propan-2-ol (also known as isopropyl alcohol)
- 20 mL 0.5 M ethanol
- 2 × Test tubes
- 3 × Capillary tubes sealed at one end
- 1 × Measuring cylinder
- 1 × Thermometer
- 3 × Rubber bands
- 1 × Stirrer
- 1 × Hot plate
- 1 × Stand with clamp

Method

1. Label each test tube 'A' and 'B'.
2. Attach the thermometer to the outside of the test tube A with the rubber band.
3. Using a measuring cylinder, add 20 mL of ethanol into test tube A.
4. Measure and pour 150 mL of liquid paraffin into the glass beaker and place the beaker on the hot plate.
5. Attach test tube A and the thermometer to the clamp.
6. Adjust the position of the clamp so that the test tube and thermometer are dipped into the beaker containing the paraffin that is on the hot plate. Ensure that the sample of ethanol is fully submerged in the paraffin but that it is not touching the bottom of the beaker.
7. Place the unsealed end of the capillary tube into the test tube.
8. Turn the hot plate on and stir the paraffin. The paraffin needs to be heated slowly.
9. Record the temperature when a continuous stream of bubbles emerges from the capillary tube (boiling point has been reached).
10. Repeat steps 3–9 for propan-2-ol using test tube B.

Results

	Temperature when a stream of bubbles occurs (°C)
Ethanol	
Propan-2-ol	

QUESTIONS

Question 1 (1 MARK)

Identify the independent and dependent variables in this experiment.

Question 2 (2 MARKS)

Identify any controlled variables in this experiment.

Question 3 (2 MARKS)

What is the purpose of stirring the paraffin liquid?

Question 4 (5 MARKS)

a Identify the trend in the data collected in your experiment. (1 MARK)

The experiment was carried out by a number of different students and the following results were obtained:

	Boiling point (°C)			
Compound	Student 1	Student 2	Student 3	Average
Ethanol	76	82	77	78.3
Propan-2-ol	80	86	81	82.3

b Explain the reason for taking the average of multiple measurements rather than recording a single set of measurements. (1 MARK)

c Which molecule has the greatest melting point? Justify your answer with reference to any relevant theory and the data collected in the experiment. (3 MARKS)

Question 5 (2 MARKS)

Write a valid aim for this experiment.

ANSWERS

1 Independent variable: The type of compound used
Dependent variable: The boiling point of the compound

2 Controlled variables:

- Volume of compound used
- Volume of paraffin liquid
- Amount of stirring
- Identity of thermometer
- Beaker in which the experiment is performed

3 [The paraffin is stirred throughout the experiment to ensure that the temperature of the liquid paraffin is consistent throughout.[1]] [This improves the accuracy of the experiment because it avoids certain parts of the paraffin/organic compound having a different temperature to the rest of the liquid (e.g. without stirring, the paraffin nearest to the hot plate will have a greater temperature than the paraffin closer to the surface).[2]]

✓ ✗ I have explained the reason behind stirring.[1]

✓ ✗ I have justified my answer with reference to the accuracy of the experiment.[2]

4 **a** *Please check your answer with your teacher.

b Repeating an experiment multiple times and averaging the results reduces the impact of any random errors in particular trials.

c [The results indicate that propan-2-ol has a higher boiling point than ethanol.[1]][Boiling point depends on the strength of intermolecular forces present in a compound.[2]][Both propan-2-ol and ethanol have a polar O-H bond and therefore, the intermolecular forces present are dispersion forces, permanent dipole-dipole forces and hydrogen bonds.[3]]

[Even though both propan-2-ol and ethanol have dispersion forces and hydrogen bonds of a similar strength, propan-2-ol is larger than ethanol which means that propan-2-ol will have the strongest dispersion forces.[4]][Therefore, since propan-2-ol has a greater strength of intermolecular forces, it is expected that it has a higher boiling point than ethanol which the results show.[5]]

- ✓ ✗ I have identified the molecule with the greatest boiling point based on the experimental data.[1]
- ✓ ✗ I have identified that the boiling point of a molecule depends on the intermolecular forces present.[2]
- ✓ ✗ I have described the intermolecular forces present in ethanol and propan-2-ol.[3]
- ✓ ✗ I have compared the relative strength of dispersion forces present in ethanol and propan-2-ol.[4]
- ✓ ✗ I have justified my answer with reference to the relationship between boiling point and intermolecular forces.[5]

5 [The aim of this experiment is to determine the effect of intermolecular bonds on a compound's boiling point.[1]]

- ✓ ✗ I have identified the objective of the experiment.[1]

CHAPTER 5 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Question 1 (1 MARK)

Which of the following chemical species exhibits covalent bonding?

A $AgNO_{3(s)}$
B $Ag_{(s)}$
C $C_2H_{6(l)}$
D $NaCl_{(s)}$

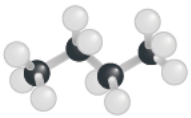
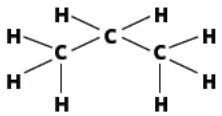
Question 2 (1 MARK)

Which of the following statements regarding covalent bonding is true?

A Covalent bonding results from the electrostatic attraction between cations and anions.
B Electrons are transferred between atoms in covalent bonds.
C Electrons are evenly distributed between atoms in all covalent bonds.
D The more electrons shared in a covalent bond, the stronger the bond.

Question 3 (1 MARK)

Which of the following is a space-filling molecular representation of propane (C_3H_8)?

A 
B 
C 
D 

Question 4 (1 MARK)

Which of the following statements regarding molecular representations is correct?

A Electron dot diagrams represent the bonds between atoms with lines.
B Structural formulas indicate the valence electrons of each atom.
C The ball-and-stick model is the most accurate three-dimensional molecular representation.
D Valence structures replace pairs of bonding electrons with lines.

Question 5 (1 MARK)

Identify the option which incorrectly matches a molecule to its shape according to VSEPR?

	Molecule	Shape
A	NF_3	Trigonal pyramidal
B	CO_2	Linear
C	C_2H_2	Bent
D	NH_4^+	Tetrahedral

Question 6 (1 MARK)

The following molecules all contain polar covalent bonds. Which of them have a permanent dipole?

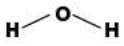
I CCl_4 **II** H_2O **III** NH_3 **IV** BeF_2

A I, II, II, IV
B I, IV
C I, II, III
D II, III

Question 7 (1 MARK)

Which molecule has the greatest percentage of intermolecular dispersion forces?

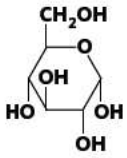
A



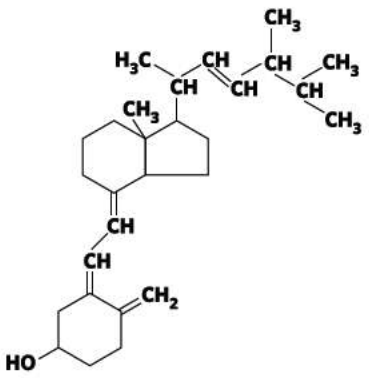
B



C



D



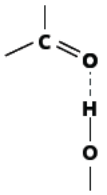
Question 8 (1 MARK)

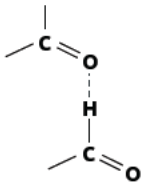
Which of the following statements is not correct?

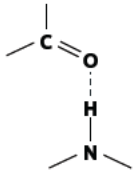
A Permanent dipole-dipole forces vary in strength depending on the difference in electronegativity of atoms present.
B Hydrogen bonds are a type of permanent dipole-dipole force.
C Permanent dipole-dipole forces occur whenever a polar covalent bond is present.
D In both dispersion forces and permanent dipole-dipole forces, there is partial electrostatic attraction between molecules.

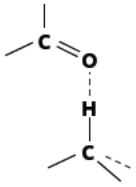
Question 9 (1 MARK)

Which of the following represents a hydrogen bond?

1 

2 

3 

4 

A 1, 2, 3, 4
B 1 and 2
C 1, 2, and 3
D 1 and 3

Question 10 (1 MARK)

Which of the following lists the chemical species in order of decreasing boiling point?

A H_2O, CH_4, HCl, H_2
B H_2O, HCl, CH_4, H_2
C H_2, HCl, CH_4, H_2O
D H_2, CH_4, HCl, H_2O

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (13 MARKS)

Hydrazine (N_2H_4) is an inorganic compound that is used in the space industry as a rocket propellant and is produced by the following reaction:

$N_{2(g)} + 2H_{2(g)} \rightarrow N_2H_{4(l)}$

a Draw the electron dot diagrams of nitrogen and hydrogen. (2 MARKS)

b How many molecules of $N_{2(g)}$ and $H_{2(g)}$ are required to produce two molecules of $N_2H_{4(l)}$? Justify your answer. (2 MARKS)

c Draw the electron dot diagram and valence structure of $N_2H_{4(l)}$. (2 MARKS)

d Based on the valence structure of $N_2H_{4(l)}$ and with reference to VSEPR, explain the shape of the molecule. It may help to use a drawing. (3 MARKS)

e Is this molecule polar? Justify your answer. (2 MARKS)

f A student wanted to compare the boiling points of C_2H_4 and N_2H_4. Which molecule would have the greatest boiling point? Explain your answer. (2 MARKS)

Question 12 (9 MARKS)

Hydrogen gas (H_2), oxygen gas (O_2) and nitrogen gas (N_2) are all diatomic molecules.

a Draw the valence structures of H_2, O_2 and N_2. (3 MARKS)

b Are these molecules polar or nonpolar? (1 MARK)

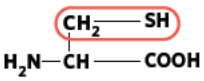
c With reference to the relative strengths of different types of covalent bonds, order the molecules in increasing order in terms of energy required to decompose the molecule. (3 MARKS)

d In the atmosphere, hydrogen gas, oxygen gas and nitrogen gas exist in fixed ratios. What type of intermolecular forces exist in a mixture of these gases? (2 MARKS)

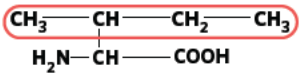
Question 13 (8 MARKS)

Proteins are important biomolecules that are made up of smaller molecules called amino acids bonded together. The forces between the side chains of amino acids are important for the stability of proteins. Three amino acids are shown.

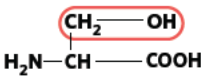
Cysteine



Isoleucine



Serine



a What type of interactions would exist between serine side chains? Draw the strongest of these interactions. (2 MARKS)

b What type of interactions would exist between isoleucine side chains? (2 MARKS)

c Cysteine can form covalent bonds in the form of -S-S- between its side chains. Explain why this is important for the structure of proteins. (2 MARKS)

d Proteins lose their structure at higher temperatures. With reference to factors affecting bond strength, explain why this occurs. (2 MARKS)

KEY SCIENCE SKILLS QUESTIONS (10 MARKS)

Question 14 (10 MARKS)

Samantha is currently undertaking a summer internship at a chemical engineering firm. She is investigating the differences between different forms of intermolecular bonding.

In order to determine the differences in bond strength, she proposes an experiment where she heats samples of $C_3H_{8(l)}$, $C_3H_7OH_{(l)}$ and $C_3H_7Cl_{(l)}$ until they boil.

a Write a simple method for this experiment and remember to specify equipment. (3 MARKS)

b Identify the independent, dependent and controlled variables. (2 MARKS)

c Explain how the data collected from the selected measurement device could be used to determine the strength of intermolecular forces between the different compounds. (3 MARKS)

d As part of her internship, the company has a duty of care to Samantha and so conducts a thorough occupational health and safety review of her experiment. Identify possible safety hazards and how to minimise them. (2 MARKS)

UNIT 1 AOS 2, CHAPTER 6

Organic chemistry

06

6A Carbon materials

6B Introduction to hydrocarbons

6C Naming simple hydrocarbons

6D Naming complex hydrocarbons

6E Functional groups - alcohols and carboxylic acids

6F Functional groups - esters

6G Polymers

6H Modifying polymers

Key knowledge

- the structure and bonding of diamond and graphite that explain their properties (including heat and electrical conductivity and hardness) and their suitability for diverse applications
- the structures, properties and applications of carbon nanomaterials including graphene and fullerenes
- the origin of crude oil and its use as a source of hydrocarbon raw materials
- the grouping of hydrocarbon compounds into families (alkanes, alkenes, alkynes, alcohols, carboxylic acids and non-branched esters) based upon similarities in their physical and chemical properties including general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties
- the formation of polymers from monomers including addition polymerisation of alkenes
- the distinction between linear (thermoplastic) and cross-linked (thermosetting) polymers with reference to structure, bonding and properties including capacity to be recycled
- the features of linear polymers designed for a particular purpose including the selection of a suitable monomer (structure and properties), chain length, degree of branching, percentage crystalline areas and addition of plasticisers
- the advantages and disadvantages of the use of polymer materials

Image: Vink Fan/Shutterstock.com

6A CARBON MATERIALS

As we saw in the last lesson, there are many different intermolecular and intramolecular forces that arise between different molecules, each affecting how a molecule interacts with its environment. In this lesson, we will be introducing carbon allotropes, and observing the differences in their structures and bonding.

6A Carbon materials	6B Introduction to hydrocarbons	6C Naming simple hydrocarbons	6D Naming complex hydrocarbons	6E Functional groups - alcohols and carboxylic acids	6F Functional groups - esters	6G Polymers	6H Modifying polymers

Study design dot points

- the structure and bonding of diamond and graphite that explain their properties (including heat and electrical conductivity and hardness) and their suitability for diverse applications
- the structures, properties and applications of carbon nanomaterials including graphene and fullerenes

Key knowledge units

Carbon	1.2.5.1
Carbon lattice structures	1.2.5.2
Carbon nanomaterials	1.2.6.1

Key terms and definitions

Allotrope one of the different physical forms an element can exist in due to different structural arrangements of atoms

Lattice regular three-dimensional arrangement of atoms

Network atoms bonded in a continuing arrangement extending outwards

Covalent network lattice regular three-dimensional arrangement of atoms covalently bonded together extending outwards

Covalent layer lattice atoms covalently bonded together in a two-dimensional arrangement to form layers, held together by weak dispersion forces

Carbon 1.2.5.1

OVERVIEW

Carbon is a very versatile element.

THEORY DETAILS

Carbon is a very common element found across the universe, and carries out a wide array of functions. The uses of carbon range from serving as the backbone of all known life on Earth, to interactions in stars at the far reaches of the universe. It is for this reason that carbon is sometimes referred to as the 'King of the Elements' of the periodic table.

Some of the key facts about carbon include:

- Carbon is a non-metal.
- It has three naturally occurring isotopes ^{12}C, ^{13}C, ^{14}C.
- It is the building block of life due to its ability to form complex, stable molecules.

One of the key characteristics of carbon is that it is able to form covalent bonds with other carbon atoms to create a variety of different structures. There are a number of pure forms in which carbon can exist, including graphite, diamond, fullerenes and graphene. These are structures that contain only carbon, however, they are arranged in different spatial orientations. Therefore, they exhibit different properties.

These different ways of structuring carbon molecules are referred to as carbon **allotropes**, and will serve as the basis of this lesson. An allotrope can exist when an element, such as carbon, has different structural arrangements due to different types of bonds between the atoms.

Let's look at an example involving another element, oxygen. Oxygen gas, O_2, is a diatomic molecule which has two oxygen atoms covalently bonded together with one double covalent bond. Ozone, O_3, is a triatomic molecule which has three oxygen atoms covalently bonded together through one double covalent bond, and one single covalent bond. Their different structures are shown in figure 1.



Figure 1 O_2 and O_3 are allotropes of oxygen.

As a result of these different structural arrangements of atoms of the same element, allotropes have different physical properties. For example, diatomic oxygen (O_2) is odourless while ozone (O_3) has a strong, distinctive smell. Ozone is also much more reactive than diatomic oxygen.

Carbon lattice structures 1.2.5.2

OVERVIEW

Carbon atoms can covalently bond with each other in a lattice-like structure to form different substances called allotropes.

THEORY DETAILS

There are many different substances that are purely carbon based. We will explore these substances throughout the rest of this lesson.

Diamond

Perhaps one of the most well known examples of a pure carbon compound is diamond. Diamond is made up of repeating units of carbon bonded to four other carbon atoms via single covalent bonds and organised into a **lattice** structure as shown in figure 2. This lattice structure also forms a **network** which is a continuous arrangement of atoms, extending outwards. Therefore, diamond is also called a **covalent network lattice**. Since every carbon atom is covalently bonded to four other carbon atoms, this results in a rigid tetrahedral network, with each carbon atom the same distance away from the next. As each carbon atom is held together by strong covalent bonds, this results in diamond's property of being a tough and rigid structure.

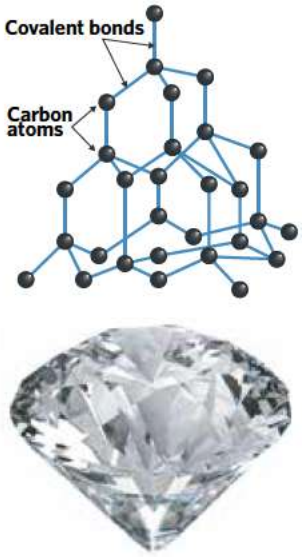


Image: magicoverlay/Shutterstock.com

Figure 2 Different representations of diamond.

Properties of diamonds

As we know, diamond is an incredibly strong substance. In fact, the only substance that is able to scratch a diamond is another diamond. These properties, as well as others, are highlighted in table 1.

Table 1 Properties of diamond.

Property	Explanation
High melting point	A number of strong covalent bonds have to be broken in order to break the lattice. Therefore it is difficult to disrupt the structure of diamond. This results in a high melting point.
Hard	The rigid lattice structure of diamond makes diamond hard.
Brittle (substance that is hard but is able to break easily)	The rigid lattice structure does not allow diamond to be bent. It will just break.
No electrical conductivity (flow of electricity through the movement of charged particles)	Each carbon takes part in four covalent bonds. This means that there are no free electrons available to conduct electricity and therefore, there is no movement of charged particles.
High thermal conductivity (ability to transfer heat)	The strong covalent bonds between the carbon atoms allow diamond to conduct heat.
Insoluble (unable to be dissolved)	The strength of the covalent bonds in the structure of diamond cannot be overcome by the intermolecular bonds of a solvent like water.

As we can see, the strength of diamond is derived from the strong covalent bonds between carbon atoms and its lattice structure.

Applications of diamond

Due to its unique properties, diamond is used for various applications. Besides being used as a jewel, diamond has industrial uses such as in cutting tools shown in figure 3. Other applications of diamond are detailed in table 2.

Table 2 Applications of diamond.

Application	Property that makes it suitable for this use.
Cutting tools	The hardness of diamond resists wear, and enhances durability.
Thermal conductor in electrical components	Diamond's strong covalent bonding enables thermal conductivity.
Optical components	Diamond lasers can be produced due to diamond's ability to transmit heat and light very effectively.
Abrasive	Diamond is able to induce friction on other objects without experiencing wear itself due to its hardness.



Image: ANDRIY B/Shutterstock.com

Figure 3 Diamond-coated dental tools.

Worked example 1

Joe is a mechanical engineer, and he suggests that a circular saw with a diamond blade should be used for cutting concrete. What property of diamond would lead Joe to make this suggestion? Explain why diamond possesses this property.

What information is presented in the question?

It has been suggested that a diamond blade should be used for cutting concrete.

What is the question asking us to do?

Determine why a diamond blade would be optimal for this task.

What strategy(ies) do we need in order to answer the question?

1. Recall the properties of diamond.
2. Apply the properties of diamond to this application.
3. Recall the structure of diamond, and use it to explain why diamond possesses these properties.

Answer

Diamond is a very hard, rigid and durable substance, which would make it ideal for use on a circular saw, as it would not easily wear.

Diamond possesses these properties due to its structure. Diamond is a covalent network lattice with carbon atoms covalently bonded to four other carbon atoms, in a network lattice arrangement that extends outwards. Therefore, this lattice structure means that carbon atoms are held together by strong covalent bonds, making diamond hard and rigid.

Graphite

Graphite, like diamond, is an allotrope of carbon, however its structure is very different to diamond's. Graphite's structure consists of layers of two-dimensional carbon lattices held together by weak dispersion forces, as seen in figure 4. This means that in each layer, carbon atoms are covalently bonded to other carbon atoms, much like diamond. However, the key difference is that there is no covalent bonding taking place between the layers and each carbon atom is only covalently bonded to three other carbon atoms. Therefore it is referred to as a **covalent layer lattice**. It is also important to understand that there is a substantial distance between the layers, approximately 2.5 times the distance between the carbon atoms in an individual layer. This is because the forces between the layers of graphite are far weaker than those within the layers.

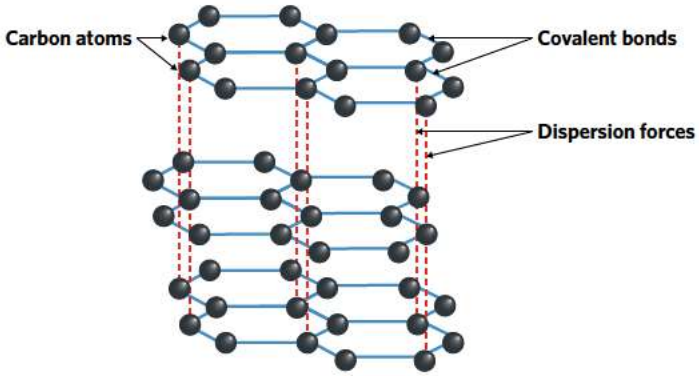


Figure 4 Structure of graphite.

One way of thinking about this is to picture a stack of waffles. Each waffle in figure 5 represents its own flat lattice of carbon atoms. However the waffles, while in contact with each other, are not connected in the same way vertically as they are horizontally. This is the same as in graphite.



Image: 5 second Studio/Shutterstock.com

Figure 5 Stack of waffles.

In graphite, each carbon atom is bonded to three other carbon atoms. Considering that carbon contains four valence electrons, this means that there is a single free unpaired electron per carbon atom. These electrons become delocalised and are able to move throughout the layers of carbon lattice.

Properties of graphite

The structural differences between graphite and diamond cause these allotropes to have vastly different properties. For example, unlike diamond, graphite is able to conduct electricity. This is due to the ability of delocalised electrons to move through the structure when an electric current is applied as seen in figure 6.

The other properties of graphite are summarised in table 3.

Table 3 Properties of graphite.

Property	Description
High melting point	A number of strong covalent bonds have to be broken in order to disrupt the structure of graphite. This results in a high melting point.
High thermal conductivity	When heat is applied, the delocalised electrons are able to move quickly and bump into other electrons in the lattice structure, thereby passing on the thermal energy from an area of high temperature to low temperature.
Soft, slippery feeling	The layers of graphite readily slide over each other due to the lack of covalent bonds.
Lower density than diamond	This is due to the greater distance between the layers of graphite, which are held together by intermolecular forces far weaker than the covalent bonding in the diamond lattice.
Insoluble	The strength of the covalent bonds in the structure of graphite cannot be overcome by the intermolecular bonds of a solvent like water.

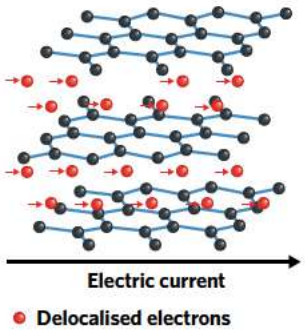


Figure 6 Movement of delocalised electrons found in graphite.

Applications of graphite

One of the most common uses for graphite is in pencils. Due to the fact that the layers of graphite are held together by weak dispersion forces, the layers of graphite are able to slide past each other with minimal force as shown in figure 7. This results in layers of graphite being deposited on the surface of paper.

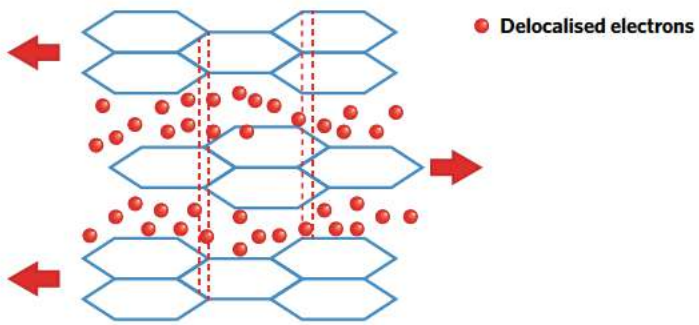


Figure 7 Movement of carbon layers within graphite.

Other applications of graphite are detailed in table 4.

Table 4 Applications of graphite.

Property	Description
Carbon brushes in electrical motors	Graphite is able to conduct electricity, and therefore can transfer current from a stationary wire to moving parts in electrical motors.
Electrode in batteries	Graphite is an inert (unreactive) and electrically conductive material.
Industrial lubricant	Layers of graphite are able to slide over each other, which, when used as a lubricant, reduces friction in machinery.

Carbon nanomaterials 1.2.6.1

OVERVIEW

Carbon-based nanomaterials have a range of different uses.

THEORY DETAILS

As we have learned in lesson 2A, nanomaterials exist at the nanoscale. Depending on the type of atoms which they are derived from, they can have a vast array of applications. Carbon nanomaterials have generated a lot of interest due to their unique properties.

Graphene

Like graphite and diamond, graphene is a carbon allotrope. Its structure is a two-dimensional hexagonal carbon lattice, as shown in figure 8, with each carbon atom covalently bonded to three other carbon atoms. In fact, structurally, graphene is a single layer of graphite. Graphene was first produced through the use of sticky tape. As the strength of the bond between the sticky tape on the top layer of graphite was stronger than that between the layers, peeling the sticky tape enabled scientists to isolate a single layer of graphene for the first time in 2004.

It is also worth acknowledging that, since many pencils use graphite, when the pencil is used graphite is left on the page. However, if the trail is made lightly enough, the user will have unknowingly produced graphene.

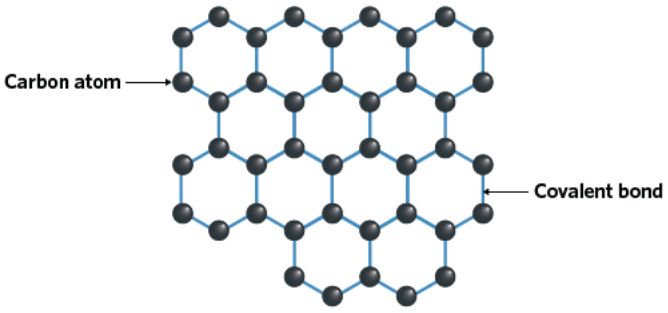


Figure 8 Structure of graphene.

Similar to the other carbon allotropes, the carbon atoms in graphene are covalently bonded together, and therefore are very strong substances. Since each carbon is bonded to only three other carbon atoms, graphene also contains delocalised electrons just like graphite.

Properties of graphene

As graphene is a single layer of graphite, it is no surprise that the properties of the two substances are similar in several ways. One of the key characteristics that is most promising about graphene is its ability to conduct electricity. The delocalised electrons in graphene are thought to be able to move incredibly fast when a voltage is applied, making it very useful in the transistors of electronic products. However, it has been shown that graphene is a much stronger material than graphite. Other defining characteristics of graphene can be seen in table 5.

Table 5 Properties of graphene.

Property	Description
Thin and light	Graphene consists of a single-layered network of carbon atoms. One square metre of graphene weighs 0.77 milligrams.
High thermal conductivity	Similar to graphite, the movement of delocalised electrons increases as heat energy is applied. This results in electrons moving quicker, bumping into other electrons and thereby conducting thermal energy.
High melting point	The strong covalent bonds holding the lattice structure together require a large amount of energy to break.
Electrical conductivity	Graphene possesses a layer of delocalised electrons. The movement of these electrons when electrical energy is applied allows graphene to conduct electricity.
Flexible	Since graphene is very thin, it is easy to bend. Additionally, because of the strength of the covalent bonds within graphene, it is difficult to break graphene through bending. Therefore, graphene is very flexible.

Uses

Graphene is often referred to as a 'wonder material' due to its unique properties. As mentioned before, graphene is very strong, approximately 200 times stronger than steel. This, as well as its ability to effectively conduct electricity, and bend without being broken, prove graphene to be an exceptionally versatile material.

Some of the proposed uses of graphene are detailed in the table below:

Table 6 Applications of graphene.

Application	Property that makes it suitable for this use
Solar cells	Graphene's flexibility could produce a breakthrough in the solar industry, with solar cells being able to be applied to any surface, including curved ones.
Filtration	Graphene membranes have been shown to be an effective filter.
Drug delivery in medicine	Graphene-based carriers of drugs for cancer therapy have been shown to target cancer cells more effectively.

Fullerenes

The term fullerenes refers to a number of different carbon allotropes which are arranged in a closed cage-like structure or a cylinder. Each carbon atom in these structures is bonded to three other carbon atoms. Some examples of fullerenes include nanotubes and buckyballs.

Nanotubes

Nanotubes are essentially a sheet of graphene rolled into a cylinder. Since they form a cylindrical tube, this allotrope is referred to as a fullerene.

Nanotubes can consist of single or multiple layers of graphene. Since the structures of graphene and nanotubes are very similar, these substances will also possess similar properties.

The sheets of graphene are folded to produce nanotubes as seen in figure 9, which is used in industries such as nanotechnology and electronic products.

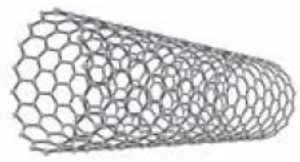


Image: Evannovostro/Shutterstock.com

Figure 9 Structure of nanotubes.

Properties

The different spatial orientation of carbon atoms in carbon nanotubes creates some differences compared to graphene. These are shown in table 7.

Table 7 Properties of carbon nanotubes.

Property	Description
High electrical conductivity	Each carbon atom is bound to three other carbon atoms, leaving one electron unoccupied for each carbon atom. Since a carbon nanotube is a cylinder, these electrons can travel along the cylinder and, thus, conduct electricity.
High thermal conductivity	Carbon nanotubes have been found to effectively conduct thermal energy.
Strong	The cylindrical structure of the tubes, together with the strong covalent bonds between the carbon atoms, result in a very strong structure.

Uses

Carbon nanotubes also have a number of unique uses which are detailed in table 8.

Table 8 Applications of carbon nanotubes.

Application	Property that makes it suitable for this use
Antennas for radios and other electromagnetic devices	The electrical conductivity and strength of carbon nanomaterials enable its use as an antenna. Additionally, the small size of carbon nanotubes makes it very space-efficient.
Fibres and fabrics	The strength created by carbon nanotubes in conjunction with their small size makes them ideal for strengthening fibres and fabrics.

These are just some of the many potential uses of carbon nanotubes. This is still a material in development, and research continues to be undertaken to discover the full potential of it.

Worked example 2

Diamond is referred to as an insulator, as it is unable to conduct electricity. However, other carbon allotropes, such as graphite and graphene, exhibit high electrical conductivity.

With reference to the structure of these materials, explain why graphite and graphene are able to conduct electricity compared to diamond.

What information is presented in the question?

Diamond is not electrically conductive, while graphite and graphene are.

What is the question asking us to do?

Provide an explanation of why graphite and graphene are able to conduct electricity while diamond can't, with reference to the important structural features enabling this.

What strategy(ies) do we need in order to answer the question?

1. Recall the structure of graphite.
2. Recall how graphene's structure relates to that of graphite.
3. Recall the structure of diamond.
4. As a result of these structures, explain why diamond is unable to conduct electricity while graphite and graphene can.

Answer

Graphite's structure consists of a series of two-dimensional carbon lattices connected by weak dispersion forces, each with a layer of delocalised electrons associated with it. This layer of electrons that is able to flow within graphite is what enables the conduction of electricity. In addition, since graphene is simply a single layer of graphite, also possessing a layer of delocalised electrons, it is also able to carry out this function.

However, diamond is a covalent network lattice and consequently, since each carbon atom is bonded to four other carbon atoms, there are no delocalised electrons and so it is unable to conduct electricity.

Buckyballs

Buckyballs, also referred to as buckminsterfullerenes, are another form of fullerene much like nanotubes. They are also a carbon allotrope, and therefore consist of only carbon atoms organised into a spherical shape. While many different types of fullerenes can be organised into spherical shapes, buckyballs are a particular type of spherical fullerene, consisting of precisely 60 carbon atoms. The structure of buckyballs is similar to the structure of a soccer ball, as shown in figure 10. The structure of these buckyballs is much like that of graphene, with each carbon atom being covalently bonded to three other carbon atoms.

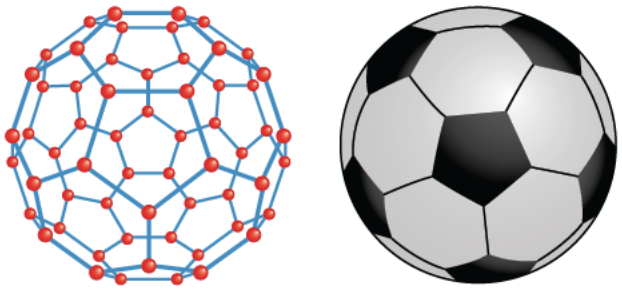


Figure 10 Structure of a buckyball.

Properties

Although their structures are unique and first thought simply to be a curiosity of science, buckyballs exhibit some similar and different properties to their allotropes, as outlined in table 9.

Table 9 Properties of buckyballs.

Property	Description
Low melting point	Weak intermolecular forces between buckyballs result in a low melting point. This is because it is hard to pack the spherically-shaped bucky balls together.
Electrical insulator and conductor	The electrical conductivity of buckyballs varies between molecules. Low density buckyballs do not conduct electricity, however C_{60} fullerenes are often used as semiconductors. Buckyballs suffer from a lack of continuity, due to their spherical shape.
Strong material	The strong covalent bonds between the carbon atoms in the fullerene molecule contribute to the strength of the molecule. Additionally, the shape in which these atoms are arranged contributes greatly to the overall strength of fullerenes.

Applications

Buckyballs have been found to have a number of different applications, some of which are outlined in table 10.

Table 10 Applications of buckyballs.

Application	Property that makes it suitable for this use
Hydrogen storage	Buckyballs can hold 8% of their mass in hydrogen at room temperature, which may be of use for the storage of hydrogen.
Solar cells	The structure of buckyballs enables them to trap electrons. Using buckyballs and nanotubes together, these fullerenes can be used as solar cells.
Antiviral and anti-cancer agent	Fullerenes are able to cover their spherical surfaces with chemotherapeutic agents, such as antiviral or anticancer agents. This enables them to transport these drugs through the body.

Theory summary

- Carbon is a highly versatile element with many different forms and functions.
- Carbon molecules can be structured in different ways, known as allotropes, each with their own unique properties.
- Carbon allotropes include diamond, graphite, graphene and fullerenes.
- The structure of diamond is a strong network of covalently bonded carbon atoms.
- Diamond is a hard, thermally conductive substance that does not conduct electricity.
- The structure of graphite consists of layers of carbon lattices held together by dispersion forces.
- Graphite's weak intermolecular forces enable layers to be rubbed off, and delocalised electrons enable the conduction of electricity.
- Graphene is equivalent to a single layer of graphite.
- Graphene is thin, flexible, strong and electrically conductive.
- Fullerenes include spherical buckyballs and tubular nanotubes.
- Nanotubes are strong and conduct heat and electricity.
- Buckyballs have a spherical shape, making them strong.

6A QUESTIONS

Theory review questions

Question 1

Carbon allotropes are

A molecules with different amounts of carbon present.

B molecules consisting of carbon atoms arranged differently.

Question 2

Diamond is

A hard and rigid.

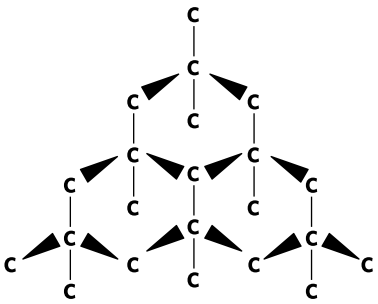
B able to conduct electricity.

Question 3

The following structure is an example of

A graphene.

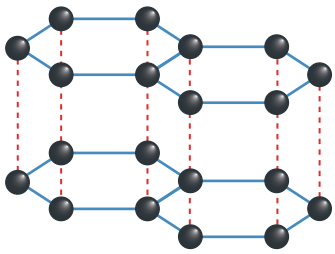
B diamond.



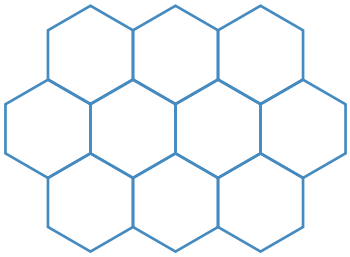
Question 4

Which of the following represents the basic structure of graphite?

A



B



Question 5

The density of graphite is

A higher than diamond's, as the lack of covalent bonds decreases the distances between layers.

B lower than diamond's, as there is a greater distance between the layers of graphite.

Question 6

Graphite is

A hard and rigid.

B able to conduct electricity.

Question 7

Graphene is most similar to

A diamond.

B graphite.

Question 8

Graphene can be rolled into

A buckyballs.

B nanotubes.

Question 9

Carbon atoms in fullerenes are bonded to

A three other carbons atoms.

B four other carbon atoms.

Question 10

Buckyballs are held together by covalent bonds.

A True

B False

Deconstructed questions

Use the following information to answer questions 11–13.

A scientist wishes to explore the different properties of diamond and graphite, and to do this, he sets up an experiment, where a sharp piece of each of these substances is rubbed against a piece of wood.

Question 11 (1 MARK)

What would be expected to happen in the case of graphite?

A A pencil mark is left on the wood.

B The graphite crumbles due to friction.

C A cut is made into the wood.

D A thin layer of wood remains on the graphite.

Question 12 (1 MARK)

Which would be more likely to happen in the case of diamond?

A A pencil mark is left on the wood.

B The diamond crumbles due to friction.

C A cut is made into the wood.

D A thin layer of wood remains on the diamond.

Question 13 (4 MARKS)

Explain the expected effect on the wood by rubbing graphite and diamond against it. Make reference to the structures of graphite and diamond, and how these contribute to their different properties.

Exam-style questions

Within lesson

Question 14 (10 MARKS)

Buckyballs, a type of fullerene, and graphene are both carbon allotropes used in the medical industry. These molecules exhibit unique properties enabling them to carry out unique functions.

a Define the term 'carbon allotrope'. (1 MARK)

b Aside from being allotropes of carbon, identify the key structural similarity between buckyballs and graphene. (1 MARK)

c Buckyballs are known to be used as a mechanism for drug delivery within the body. Describe one key feature of the shape of buckyballs that enables their function as a drug delivery molecule. (2 MARKS)

d It was proposed by a medical researcher that it may also be possible to use nanotubes for drug delivery purposes in medicine, due to the similarities these molecules have with buckyballs. These are molecules that are currently used for purposes such as antennas for radios, or to strengthen materials.

i Identify the similarities and differences between buckyballs and nanotubes. (2 MARKS)

ii Describe the structural features of nanotubes that enables their use as fabric strengtheners. (2 MARKS)

iii Describe the structural features of nanotubes that enables their use as radio antennas. (2 MARKS)

Question 15 (3 MARKS)

A sketch artist is producing a sketch of a landscape, and begins to think about the mechanisms behind how a pencil works. She develops the following theory:

'The end of a modern pencil consists of graphite, which leaves individual atoms behind as a result of friction breaking the covalent bonds between them when the pencil is rubbed against paper.'

Evaluate the accuracy of this statement.

Question 16 (6 MARKS)

Electric motors use carbon brushes in order to conduct electrical current from a stationary wire to a moving component in the motor. In modern technology, graphite is often used to construct carbon brushes.

a Describe the property of graphite enables it to be used for this purpose. (2 MARKS)

b Diamond is another material being considered for use as a carbon brush, however this is quickly ruled out. Provide two reasons why diamond would be ineffective for use as a carbon brush. (2 MARKS)

c Extremely high temperatures are able to be generated in electric motors. Would this compromise the functionality of the graphite carbon brushes? Explain your answer with reference to graphite's structure. (2 MARKS)

Question 17 (8 MARKS)

In dentistry, tools tipped with diamonds are used to drill and polish cavities. This is because the presence of diamonds on the tip enhances the abrasive property of these tools, or makes the tool more effective at shaping or polishing cavities through friction.

a Draw the atomic structure of diamond, showing bonding between atoms. (2 MARKS)

b With reference to the structure of diamond, explain why diamond is an effective abrasive. (2 MARKS)

c Compare the effectiveness of using diamond and graphite for this purpose in dentistry. (4 MARKS)

Question 18 (8 MARKS)

A startup has devised a new model of solar panels which can be integrated into cars. However, in order to achieve this, the company must establish a way to make solar panels curve with the shape of the car, so that it both looks better but also makes the car more aerodynamic.

a What carbon allotrope could be used for this purpose? Explain why, making reference to the molecular structure of this allotrope. (3 MARKS)

When enquiring about the material that should be used to produce the solar panels, the company accidentally ordered the wrong material. They found that this material also conducted electricity, and had many of the same properties as the desired carbon allotrope, however it was not as thin and flexible.

b What material did the company order in? Explain. (2 MARKS)

c What is the easiest way to produce the desired material from the material ordered in? Describe how this method works, and whether this would be feasible for the mass production of solar panels. (3 MARKS)

Multiple lessons

Question 19 (5 MARKS)

An electrical engineer is attempting to design a new form of battery, and is exploring different options for the material that could be used as an electrode. Although electrodes are generally made from metals, a colleague of his suggests that the electrodes be made out of graphene instead.

a What is similar about a metal and graphene that would enable it to be used in batteries? (2 MARKS)

b The battery to be used was going to be placed in an electric car next to the combustion engine. As a result, it would be subject to quite high temperatures. Upon further research, the engineer finds that the melting point of graphene is much higher than metals such as copper, which is commonly used in cars. Describe the implication that the information above has on the structure of metals and graphene. (3 MARKS)

Key science skills

Question 20 (3 MARKS)

A team of laboratory clinicians decide to carry out an experiment comparing the properties between four carbon allotropes – diamond, graphite, graphene and fullerenes. The team set up a series of experiments to assess the electrical and thermal conductivity of each of the allotropes.

The team of laboratory clinicians knew before conducting the experiment that most of the carbon allotropes would be good electrical conductors as well as good conductors of thermal energy. Upon analysis of the results, the results on thermal conductivity proved to be correct, however the results showed that none of the allotropes conduct electricity. The clinicians concluded that an error must have been made.

a What type of error has likely taken place? (1 MARK)

b Suggest a way of avoiding such errors. (1 MARK)

When repeating the tests to assess the results of the thermal conductivity of the substances, the researchers noticed that while all the results were close to the expected temperature, there was significant deviation between different results.

c What can be concluded about the results? (1 MARK)

Exam-style question hints

14a: Diamond, graphite, graphene and fullerenes are carbon allotropes. **14b:** Carbon atoms in carbon allotropes take part in covalent bonding. **14c:** Molecules can be used for different functions due to their structure. **14di:** Buckyballs and nanotubes interact with their environments differently. **14dii:** Reinforcing a material makes it more difficult to break. **14diii:** Antennas are essential in the transmission of electrical energy. **15:** Marks made by modern pencils are a result of layers of a carbon allotrope left on a page. **16a:** Electric currents involve the movement of electrons. **16b:** The structure of diamond is different to that of graphite. **16c:** The structure of a compound impacts the ability of the substance to withstand heat. **17a** Diamond has a strong carbon lattice. **17b:** Diamond is a strong material. **17c:** Graphite and diamond are structurally different substances. **18a** Solar panels are required to conduct electricity. **18b:** Graphene is structurally similar to another carbon allotrope. **18c:** Graphene is derived from graphite. **19a:** Batteries produce and pass on electrical energy into devices. **19b:** Melting point reflects the amount of energy required to break certain bonds. **20a:** Different types of errors affect results differently. **20b:** Changes need to be made to ensure the accuracy of the experiment. **20c:** Consider accuracy and precision.

6B INTRODUCTION TO HYDROCARBONS

In this lesson, we will be continuing our exploration of carbon-based materials by introducing hydrocarbons, crude oil, and their place in contemporary society.

6A Carbon materials	6B Introduction to hydrocarbons	6C Naming simple hydrocarbons	6D Naming complex hydrocarbons	6E Functional groups - alcohols and carboxylic acids	6F Functional groups - esters	6G Polymers	6H Modifying polymers

Study design dot points

- the grouping of hydrocarbon compounds into families (alkanes, alkenes, alkynes, alcohols, carboxylic acids and non-branched esters) based upon similarities in their physical and chemical properties including general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties
- the origin of crude oil and its use as a source of hydrocarbon raw materials

Key knowledge units

Introduction to hydrocarbons	1.2.8.1
Crude oil	1.2.7.1
Hydrocarbon bonding	1.2.8.2

Key terms and definitions

Hydrocarbon compounds consisting of hydrogen and carbon

Organic compounds compounds consisting of carbon atoms most commonly covalently bonded to hydrogen, nitrogen, phosphorus or oxygen atoms

Carbon chain linear set of carbon atoms in a hydrocarbon

Homologous series series of hydrocarbons that have a similar chemical structure and chemical properties

Crude oil unrefined petroleum deposit composed of hydrocarbons

Fossil fuel fuels derived from fossilized plant and animal matter

Renewable fuel a fuel able to be replenished at a rate greater than or equal to its rate of consumption

Non-renewable fuel a fuel consumed at a rate greater than it is replenished

Fractional distillation separating of a liquid mixture into fractions differing in boiling point

Boiling point temperature at which a liquid transitions into gas

Alkane hydrocarbon that contains only single carbon to carbon bonds and without functional groups

Alkene hydrocarbon with at least one double carbon to carbon bond

Alkyne hydrocarbon with at least one triple carbon to carbon bond

Saturated hydrocarbon a hydrocarbon possessing only single bonds between carbon atoms

Unsaturated hydrocarbon a hydrocarbon possessing at least one double or triple bond between carbon atoms

Substituents an atom or group of atoms replacing one or more hydrogens on the parent chain of a hydrocarbon

Addition reaction one molecule combining with another molecule to form a larger molecule

Substitution reaction one molecule replacing a part of another molecule

Halogenation replacement of a hydrogen atom with a halogen in a molecule

Haloalkane alkane possessing a halogen functional group; a halogenated alkane

Halogen element in Group 17 of the periodic table

Isomers compounds that share the same molecular formula but have different structures

Lesson links

This lesson builds on:

- 6A Carbon materials

Hydrocarbons are a type of carbon compound.

Introduction to hydrocarbons 1.2.8.1

OVERVIEW

Hydrocarbons are organic molecules found abundantly around the world that serve many different purposes.

THEORY DETAILS

The functioning of modern human society relies on various forms of energy, whether that be in the form of electricity to power an electrical device, or perhaps thermal energy to warm our homes. However, the use of these fuels can have detrimental environmental effects; an idea that has become a prominent point of contemporary political debate.

One primary source of energy for human consumption is **hydrocarbons**. These are **organic compounds** consisting of carbon and hydrogen atoms, organised with the carbon atoms in a linear chain-like sequence. This can be thought of like an industrial chain, hence the name **carbon chain** as shown in figure 1.



Image: tassel78/Shutterstock.com

Figure 1 Propane and chain.



The prefix 'hydro-' in hydrocarbon refers to the presence of hydrogen in the molecule.

Although the term organic is often used to describe things, such as food, as natural, this is not what the term means in chemistry. When a molecule is described as an organic molecule, this means that the molecule is carbon-based, which may be either naturally or synthetically produced.

The basic structure of a single chain hydrocarbon consists of a linear chain of carbon atoms bonded together. Hydrocarbons can vary in the number of carbon and hydrogen atoms they have and those that differ from their previous hydrocarbons in the series through the addition of a $-CH_2-$ group, are described as belonging to the same **homologous series**. As a result, hydrocarbons in the same homologous series are going to have:

- Similar physical and chemical properties
- Similar structure
- The same general formula

As we have discussed, each carbon atom contains four valence electrons and therefore to satisfy the octet rule, would need to share four electrons with other atoms. As a result of this, each carbon atom can form four covalent bonds. As shown in figure 2, the carbon atoms found in hydrocarbons can form covalent bonds with different types of atoms to become stable.

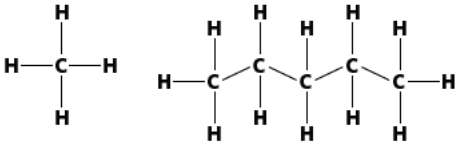


Figure 2 Structural formula of methane (left) and pentane (right).

As we will see later in this chapter, carbon can bond to a multitude of different non metals to gain a full valence shell.

Uses of hydrocarbons

Hydrocarbons are used to carry out many functions in society. For example, hydrocarbons can be used to produce asphalt for laying roads, as an industrial lubricant, and most importantly as a fuel source.

Oil and gas fuels, including petrol, diesel and kerosene, are hydrocarbon-based, and are used to power society. When hydrocarbons undergo combustion reactions, they release energy which is able to be used in many different situations, such as cooking on a stove or even supplying the grid with electricity. This is demonstrated in the following reaction:

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)} \quad \Delta H = -890 \text{ kJ mol}^{-1}$$

Methane (CH_4) can take part in a complete combustion reaction with oxygen to form carbon dioxide and water to release energy. This reaction is referred to as complete combustion due to the large amount of oxygen available, helping the reaction take place, as well as the production of carbon dioxide. This reaction also leads to the release of 890 kilojoules of energy per mol of methane consumed.

As thermal energy, heat, is released through the combustion of hydrocarbons, we can transform this energy into a form that we can use. For example, thermal energy can be transformed into mechanical energy by spinning a turbine, which can be converted to electrical energy, however this topic will be covered in greater detail later in your studies of chemistry.

Crude oil 1.2.7.1

OVERVIEW

Crude oil is a source of many different types of hydrocarbons used as a fuel source.

THEORY DETAILS

Crude oil is a naturally occurring **fossil fuel**, which means that it is produced from the fossilisation process of plant and animal matter over the course of millions of years. Crude oil refers to what we may know as oil or petroleum, and substances such as petrol, diesel and kerosene are derived from it. It has been used as a fuel source since its discovery in the 1850s, and is produced beneath the earth's surface in crude oil basins. The distribution of different sources of energy used for human consumption on Earth is demonstrated in figure 3, where it can be seen that crude oil makes up 33% of global energy supply. What makes crude oil so useful is that it consists primarily of hydrocarbons, able to be combusted to release energy. This oil is extracted through the boring mechanisms of oil well as can be seen in figure 4, which forces crude oil to the surface before being transported for processing into usable forms of energy.

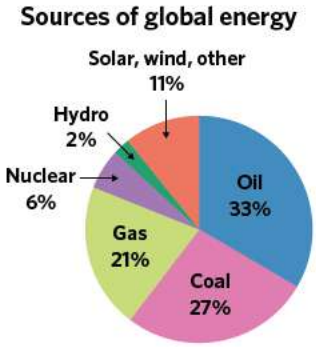


Figure 3 Graph showing sources of global energy.



Image: Xiangli Li/Shutterstock.com

Figure 4 Oil well extracting crude oil.

This fossilisation process takes place due to the exposure of these substances to high temperatures and pressures over millions of years as the organic matter gets buried under sand and mud, reducing the material from complex organic molecules from lifeforms into the basic structures of hydrocarbons. Figure 5 demonstrates the location where crude oil is found in the Earth. As a general rule, the higher the temperature the organic matter is exposed to, the lighter the resulting hydrocarbons will be, as the increased thermal energy is able to break down the molecule further.

It is worthwhile considering that some of this organic matter dates back even so far as to the Devonian Period, ranging from 419.2 million to 358.9 million years old and now being used to power our society!

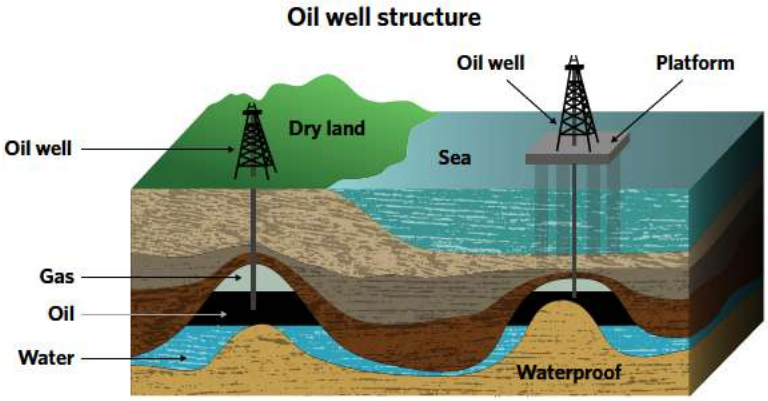


Image: Ellen Bronstayn/Shutterstock.com

Figure 5 Oil well extracting oil.

It is also important to introduce the concept of renewability. A fuel is considered to be **renewable** if it can be replenished at a rate greater than or equal to its rate of consumption. Some examples of renewable fuels are biodiesel, biogas and bioethanol, which are all produced from plant matter and thus quick to produce. However, because these fossil fuels take such a long time to develop, and are able to be consumed in a relatively short amount of time, these fuels are considered as **non-renewable**. This implies that if non-renewable fuels are used, the fuel sources will eventually be depleted as it cannot be replenished quickly enough.

It should also be noted that when the combustion of crude oil takes place, carbon dioxide is released into the atmosphere, which is a greenhouse gas. The accumulation of greenhouse gases such as carbon dioxide in the atmosphere can trap heat in the earth's atmosphere, causing drastic changes in atmospheric and ecological conditions known as climate change. Some fuels counteract this effect by consuming atmospheric carbon dioxide in order to photosynthesise and produce energy, such as those produced from plants. Although crude oil can derive from plant matter, this has been underground and non-photosynthetic for millions of years, and therefore not consuming carbon dioxide from the atmosphere.

What is crude oil made of?

Crude oil consists primarily of a mix of different hydrocarbons, as well as small amounts of nitrogen, sulfur, and oxygen. However, in most cases, carbon makes up approximately 82–87% of crude oil's weight, and hydrogen 12–15%.

In order to separate these hydrocarbons based on chain length, a process called **fractional distillation** is used. Fractional distillation utilises the different **boiling points** of these different hydrocarbons to separate them.

As seen in figure 6, the crude oil is separated into its constituent parts so that the optimal hydrocarbons can be used as, for example, fuel oil as opposed to kerosene.

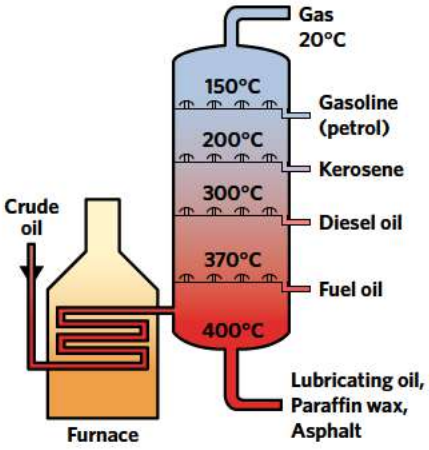


Figure 6 Fractional distillation of crude oil.

Uses of crude oils

In the process of fractional distillation, as seen in figure 6, the crude oil is heated in a furnace until all the hydrocarbons in the mixture have been converted to a gaseous form. Following this, the gases enter a column with its highest temperature at the bottom, and lowest temperature at the top. Due to the presence of catches at each level, when the temperature drops to the boiling point of a hydrocarbon, the crude oil will be converted from a gas to a liquid, and land in the catch. As a result of this, the molecules become separated based on their boiling point, which is indicative of their further properties.

As we have studied previously in lesson 5C, the longer the hydrocarbon's carbon chain length, the higher the boiling point will be for that molecule as there are more dispersion forces between the molecules. This means that more energy is required to overcome these forces to transition to a gaseous state.

The hydrocarbons with a higher boiling point are generally utilised for higher intensity purposes for example asphalt for the laying of road, or tough solid lubricants for heavy machinery. These hydrocarbons, while flammable, will not burn easily due to the fact that their long carbon chains prevent them from releasing substantial amounts of vapour.

Hydrocarbon bonding 1.2.8.2

OVERVIEW

The bonding that occurs between carbons, and between carbon and hydrogen give rise to a diverse range of molecules.

THEORY DETAILS

Perhaps the most clear difference between hydrocarbons to notice is the number of carbon atoms in the chain, ranging from methane with 1 carbon atom, growing to sizes of molecules such as maitotoxin as shown in figure 7, which has a total of 164 carbon atoms.

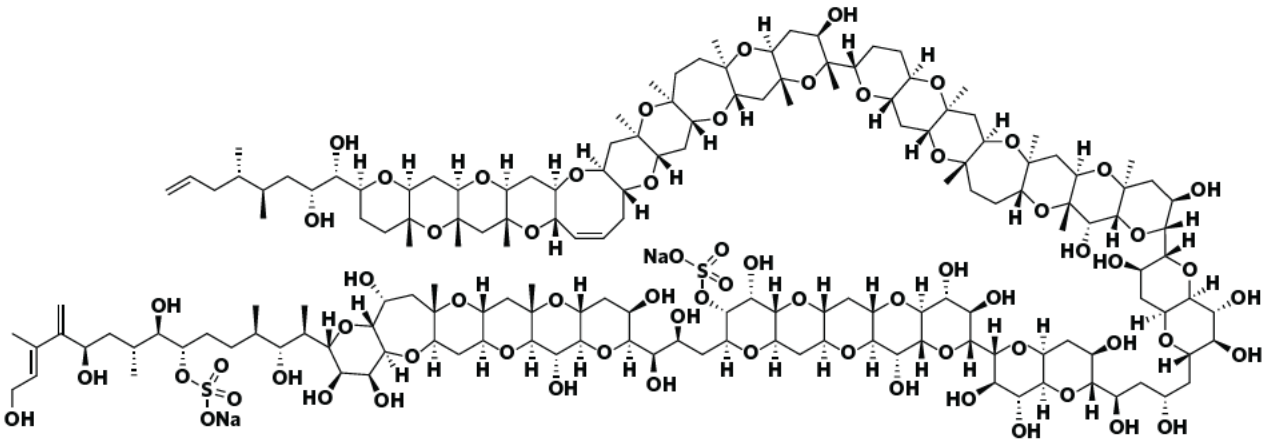


Figure 7 Maitotoxin structural formula.

While this may be the most obvious difference between hydrocarbons, these molecules do not only vary in the number of carbon atoms they have in their carbon chain, but also contain different types of bonds as well as functional groups.

Carbon atoms are able to form single, double and triple bonds with other atoms. This means that these carbon atoms can form one, two or three covalent bonds with another carbon, as shown in figure 8.



Don't be intimidated by maitotoxin! In this course, students are only required to be familiar with hydrocarbons up to 10 carbon atoms in length.

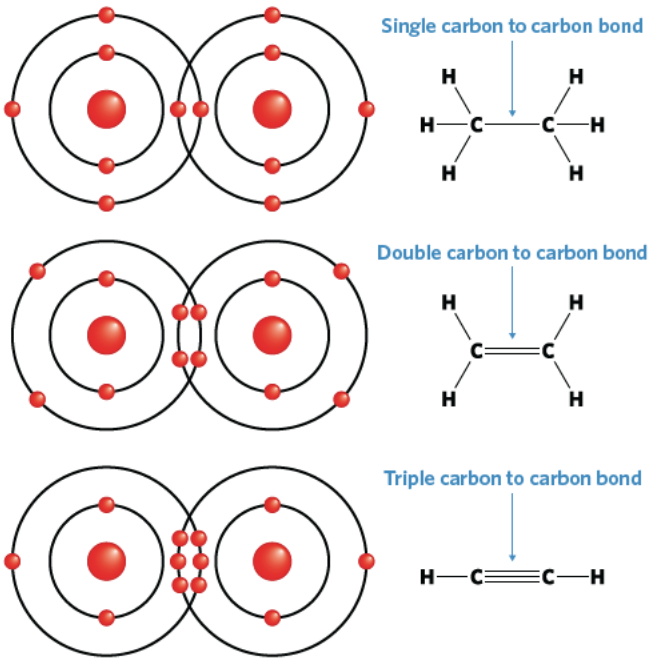
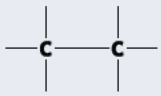
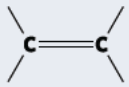


Figure 8 Types of carbon bonding.

As we learned previously, a covalent bond involves two atoms sharing an electron each to fill their shells. Therefore, in single bonds the atoms are sharing two electrons, in double bonds sharing four electrons, and sharing six in triple bonds. The hydrocarbons already presented in this lesson, propane, methane and pentane, only contain single covalent bonds as indicated by the single line between the atoms.

Different groups of hydrocarbons have different names based on the type of bonds present in their carbon chain. Hydrocarbons that only contain single carbon to carbon bonds in their carbon chain and have no functional group are referred to as **alkanes** and have the general molecular formula C_nH_{2n+2} where n represents the number of carbons in the parent chain. If there is a double bond between two of the carbon atoms then the hydrocarbon is referred to as an **alkene** and has a general formula of C_nH_{2n}. If there is a triple bond the compound is considered as an **alkyne** and has a general formula of C_nH_{2n-2}. The carbon to carbon double and triple bonds are considered as functional groups. This is summarised in table 1.

Table 1 Categorisation of different types of hydrocarbons.

Description	Representation	Type of hydrocarbon	General formula (where n represents the number of carbons in the parent chain)
Carbon chain contains only single carbon to carbon bonds		Alkane	C_nH_{2n+2}
Carbon chain contains at least one carbon to carbon double bond		Alkene	C_nH_{2n}
Carbon chain contains at least one carbon to carbon triple bond		Alkyne	C_nH_{2n-2}

Worked example 1

A hydrocarbon has a molecular formula of C_6H_{12}. Would this molecule be considered as an alkane, alkene or alkyne?

What information is presented in the question?

The molecular formula for a molecule.

What is the question asking us to do?

Classify the hydrocarbon as an alkane, alkene or alkyne.

What strategy(ies) do we need in order to answer the question?

1. Identify the three different formulas for each hydrocarbon type.
2. Determine the formula that is able to generate the formula given in the question.
3. Classify the molecule.

Answer

General formula:

Alkane: C_nH_{2n+2}

Alkene: C_nH_{2n}

Alkyne: C_nH_{2n-2}

The formula given is C_6H_{12}, therefore we can see that $n = 6$.

If the molecule was an alkane, it would have the molecular formula:

$= C_6H_{(2 \times 6)+2}$

$= C_6H_{14}$

As we can see, this formula is different to the given formula.

If the molecule was an alkene, it would have the molecular formula:

$= C_6H_{(2 \times 6)}$

$= C_6H_{12}$

This formula matches the given formula. As a result, C_6H_{12} would be an alkene.

It is worth noting that these bonds differ in their strengths, with triple bonds being the strongest and single bonds being the weakest. This is because in triple bonds, there are more covalent bonds holding the molecules together, and thus the bond strength will be greater. This is demonstrated in table 2.



By using the general formulas, you will be able to distinguish between alkanes, alkenes and alkynes without having to draw or see the whole structure.

Table 2 Relative strengths of carbon to carbon bonds.

Bond	Relative Strength
C–C	345 kJ/mol
C=C	611 kJ/mol
C≡C	837 kJ/mol

Another way in which we can describe hydrocarbons based on their bonding patterns is through the idea of saturation. **Saturated hydrocarbons** are those that only have single bonds between the carbon atoms, whereas **unsaturated hydrocarbons** have at least one double or triple bond between carbon atoms in a carbon chain.

Hydrocarbons can further diversify through the addition of functional groups and side groups to their structure. By definition, functional groups are simply **substituents** of a molecule that are responsible for a particular molecule's unique interactions, and thus, as double and triple bonds define how hydrocarbons interact, these are also considered functional groups. Functional groups will be explored in further detail in later lessons.

Physical properties of alkanes

As hydrocarbons are a homologous series of molecules, they all have similar physical and chemical properties to each other. However, as the carbon chain length in the hydrocarbons increases or decreases, this alters the extent to which these properties are exhibited.

Alkanes, alkenes and alkynes each have their own unique physical and chemical properties, which will be covered later in this lesson. The properties of hydrocarbons change with the introduction of functional groups to their structures, which will be covered in the next lesson. For now we will be observing the physical and chemical properties of alkanes, which have only single bonds in their structures and do not have any functional groups.

Hydrocarbons without functional groups, such as alkanes, are non-polar molecules, which cannot dissolve in water. The strongest bonds able to be formed between alkanes are weak dispersion forces. Dispersion forces arise as a result of interactions between individual atoms in a molecule, and therefore as the length of a carbon chain increases, the intermolecular forces between the hydrocarbons also increases due to the stronger forces between the molecules.

As a result of this, different alkanes exhibit different properties based on their differing carbon chain lengths, which therefore makes them useful for different purposes. For example, at room temperature methane is found in a gaseous form, whereas octane is found in a liquid form, such as petrol. The differences in state of these hydrocarbons reflects the strength of the intermolecular bonds formed.

Some other physical properties of alkanes include:

- Pure alkanes are colourless, tasteless and odourless
- Alkanes are less dense than water. This in conjunction with them being non-polar explains why alkanes float on water.

Chemical properties of alkanes

Alkanes have a set of distinct chemical properties that they exhibit, which will also be distinct from the chemical properties of other types of hydrocarbons, such as alkenes or alkynes. Alkanes are primarily able to undergo **substitution reactions**, which contrasts to the ability of alkenes and alkynes to undergo **addition reactions**, which will be covered later in the lesson.

As indicated by the name, substitution reactions are those in which atoms bonded to a carbon atom are replaced by a different atom. Take the reaction shown in figure 9.

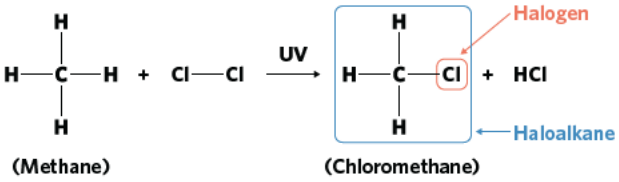


Figure 9 Substitution reaction involving an alkane.

In this reaction, one of the hydrogen atoms in methane is being substituted with a chlorine atom, resulting in the formation of chloromethane and hydrogen chloride. This demonstrates a universal chemical property of alkanes, which is their ability to undergo **halogenation**. This is a reaction in which an alkane reacts with a **halogen**, such as chlorine or bromine, in the presence of ultraviolet light, resulting in the formation of a **haloalkane**.

Additional chemical properties of alkanes include:

- Alkanes are relatively unreactive and do not participate in reactions such as acid-base reactions.
- Alkanes are combustible and release energy when they participate in combustion reactions

Isomers of hydrocarbons

Molecules which have the same molecular formula, for example C_6H_{14}, don't necessarily have the same structure. Molecules with the same molecular formula but different structure are referred to as being **isomers** of each other. For example, the molecule with the molecular formula C_6H_{14} can have different structures as shown in figure 10.

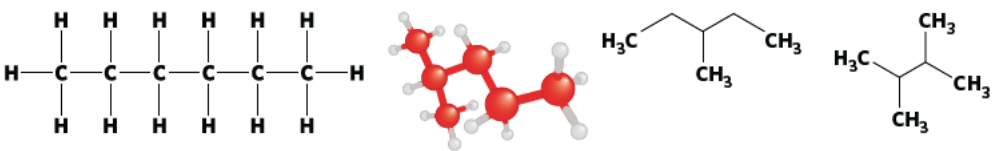


Figure 10 Structures of some isomers of C_6H_{14}.

Differences in physical and chemical properties

It is important to note that while alkanes, alkenes and alkynes are all members of the same homologous series of hydrocarbons, the differences in their structures will give rise to differences in their properties.

The physical properties of alkenes and alkynes, such as intermolecular forces, are largely comparable to those of alkanes. Alkynes have higher **boiling points** than alkanes and alkenes due to their electric field, however the details of this is beyond the scope of the course.

Alkenes and alkynes are more distinguishable from their alkane counterparts due to their differing chemical properties. Earlier in the lesson we discussed substitution reactions, wherein alkanes have a hydrogen atom in their structure replaced with another atom, such as a halogen. In the case of alkenes and alkynes, these do not participate in substitution reactions, but rather they participate in **addition reactions**. Rather than replacing an atom in the structure of the hydrocarbon, in addition reactions a molecule is added onto the hydrocarbon without the loss of any matter. This is able to take place due to the presence of double and triple bonds. When a bond in a double bond is broken, this frees up an additional electron on each carbon involved in the carbon to carbon double bond, enabling the bonding of additional atoms to its structure as seen in figure 11.

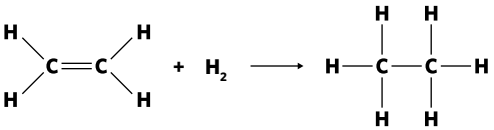


Figure 11 Addition rection between ethene and hydrogen.

As can be seen, the double bond is being broken, and two hydrogen atoms are being added to the structure of the hydrocarbon. This ability to participate in addition reactions is a property that clearly distinguishes alkenes and alkynes from alkanes.

Theory summary

- Hydrocarbons can be used as fuel sources.
- Crude oil is a type of hydrocarbon that is a fossil fuel.
- Fractional distillation is used to separate hydrocarbons from crude oil.
- Fractional distillation separates hydrocarbons based on their boiling points.
- There are three main types of hydrocarbons: alkanes, alkenes, alkynes.
- Alkanes contain single carbon to carbon bonds, alkenes contain double carbon to carbon bonds, alkynes contain triple carbon to carbon bonds.
- The general formulas for alkanes, alkenes and alkynes are C_nH_{2n+2}, C_nH_{2n} and C_nH_{2n-2} respectively, where n represents the number of carbons.
- Saturated hydrocarbons contain only single carbon to carbon bonds, whereas unsaturated hydrocarbons contain carbon to carbon double and/or triple bonds.
- Haloalkanes are hydrocarbons that contain halogens.
- Isomers are hydrocarbons that share the same molecular formula but different structure.

6B QUESTIONS

Theory review questions

Question 1

Hydrocarbons are inorganic molecules.

A True

B False

Question 2

Hydrocarbons that differ to by CH_2 are known to belong to

A the same homologous series.

B the same hydrocarbon family.

Question 3

Crude oil is considered to be

A renewable.

B non-renewable.

Question 4

Using fractional distillation, hydrocarbons with the greatest boiling point will be separated

A first.

B last.

Question 5

Alkane + chlorine → _____

A halogen

B haloalkane

Question 6

All hydrocarbons, no matter their carbon chain length, have the same properties.

A True

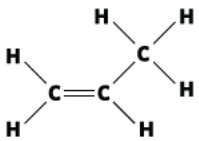
B False

Question 7

The molecule shown is considered as a

A saturated hydrocarbon.

B unsaturated hydrocarbon.

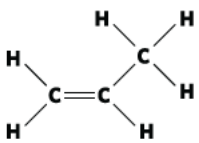


Question 8

The molecule shown is considered as an

A alkene.

B alkyne.

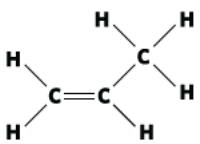


Question 9

The main force of attraction affecting the boiling point of the molecule shown is

A hydrogen bonding.

B dispersion forces.

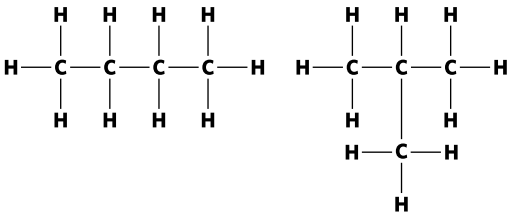


Question 10

The molecules shown are described as

A alkynes.

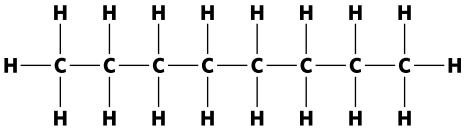
B isomers.



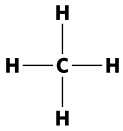
Deconstructed questions

Use the following information to answer questions 11–13.

Hydrocarbons are very useful in everyday life. For example, petrol is widely used as a transport fuel whereas fuels like methane can be used for cooking purposes. The structure of both molecules are shown.



Octane C_8H_{18}
(Found in petrol)



Methane CH_4

Question 11 (1 MARK)

Octane and methane are considered as

A alkynes.

B alkenes.

C alkanes.

D haloalkanes.

Question 12 (1 MARK)

In terms of the main intermolecular forces responsible for the boiling point of each molecule,

A octane is predominantly held together by hydrogen bonds whereas methane is predominantly held together by dispersion forces.

B methane is predominantly held together by hydrogen bonds whereas octane is predominantly held together by dispersion forces.

C both are held together by hydrogen bonds only.

D both are held together by dispersion forces only.

Question 13 (3 MARKS)

Methane is a much cheaper fuel than octane. A car company wanted to test whether or not they could build a fuel tank that was suitable for methane as part of a new, more cost effective car. One of the biggest concerns is to be able to develop a tank that is able to insulate the fuel from heat generated by the car while it is functioning. After conducting tests, they were able to determine the following boiling points of both fuel sources.

	Fuel A	**Fuel B**
Boiling point (°C)	–162	–126

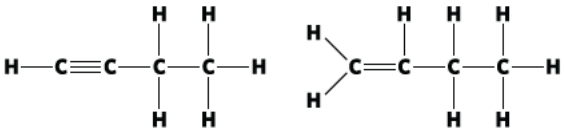
Which of the fuels, octane or methane, would be able to withstand the most heat before boiling? Based on your answer, identify the name of fuel A and B.

Exam-style questions

Within lesson

Question 14 (5 MARKS)

Consider the molecules shown.



Molecule A **Molecule B**

a Can the two molecules be considered as isomers of each other? Explain. (2 MARKS)

b A student wanted to select the appropriate molecule that would take part in an addition reaction. Suggest the appropriate molecule that would be able to meet this requirement. (2 MARKS)

c A molecule with the same number of carbon atoms as the molecules above has the formula C_4H_{10}. Classify this as an alkane, alkene or alkyne. (1 MARK)

Question 15 (6 MARKS)

A company in New South Wales supplying energy to the grid is exploring the potential of using crude oil to power the city, while minimising its impact on greenhouse emissions.

a What does the term 'renewability' mean? (1 MARK)

b Name and briefly describe the process of the extraction of fuel from crude oil. (3 MARKS)

c The company concludes that crude oil is a non-renewable fuel. Explain why the company came to this conclusion. (2 MARKS)

Multiple lessons

Question 16 (6 MARKS)

There are a large number of different types of hydrocarbons that exist. At the most basic level, hydrocarbons can be categorised into two main categories.

Complete the table below.

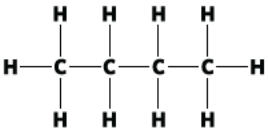
	Saturated hydrocarbons	Unsaturated hydrocarbons
Bonding type within the hydrocarbon chain		
Can participate in addition reactions		
Contains covalent bonds		

Question 17 (4 MARKS)

Fluorine gas is highly toxic and therefore quite dangerous when it is inhaled. However, when used to produce fluorocarbons, it can be used as refrigerants to keep fridges continuously cool.

a Draw the structural diagram to show the bonding of fluorine gas. (1 MARK)

b Supposed the following molecule took part in a reaction with fluorine gas. What is the name of this type of reaction? (1 MARK)



c Describe the type of bonding that would be responsible for holding this newly formed molecule in its liquid state. (2 MARKS)

Question 18 (4 MARKS)

For a road trip, Grantham wanted to figure out how much fuel he would need to make sure that he reached his destination. His new car will only take the best petrol, which consists of the hydrocarbon octane C_8H_{18}, which has a density of 0.703 g/mL.

a Prior to the trip, Grantham filled his car so that he would have a full tank of petrol. If a full tank is 40.0 L, how many grams of octane fuel did he have in total? (1 MARK)

b What amount, in mol, of octane fuel, does he have when his tank is full? (2 MARKS)

c If one mol of octane fuel is able to generate 5460 kJ of energy when burnt, how much energy would a full tank of fuel generate? (1 MARK)

Question 19 (10 MARKS)

A chemistry student is investigating the chemical structures present in bonding between molecules.

a The student observes a carbon-carbon double bond. How many electrons are being shared between the two carbon atoms? Justify your answer. (3 MARKS)

b What is the strongest possible bond type between carbon atoms in a hydrocarbon chain? Explain. (2 MARKS)

c What type of carbon to carbon bonds in a hydrocarbon chain is the least reactive with bromine gas? Justify your answer. (3 MARKS)

d Compare and contrast the strength of a hydrocarbon with that of diamond. Explain your reasoning. (3 MARKS)

Key science skills

Question 20 (6 MARKS)

A scientist on an oil rig has been called to investigate an oil spill that has taken place in the Bass Strait, south of Melbourne. Having noticed oil floating on the water, workers called in the scientist to determine the composition of the substance and the severity of the spill.

a Suggest the types of compounds likely to be found in the oil spill. (2 MARKS)

b Suggest a reason why the oil does not mix with the water. (2 MARKS)

When measuring the amount of oil in the water, a worker accidentally knocks a barrel of oil over, with its contents all spilling into the oil spill.

c What type of error has taken place? (1 MARK)

d Suggest a way of remedying this error. (1 MARK)

Exam-style question hints

14a: Isomers share a common characteristic. **14b:** Addition reactions require unsaturated molecules. **14c:** The general formulas give information about the nature of the hydrocarbon. **15a:** Renewability relates to rate of fuel production. **15b:** Oil contains many different types of compounds that are separated based on their structures. **15c:** Renewability relates to rate of fuel production. **16:** The three types of hydrocarbons can be classified into saturated and unsaturated hydrocarbons. **17a:** Structural diagrams only contain bonding pairs of electrons. **17b:** Fluorine is a halogen. **17c:** Intermolecular bonding depends on the nature of the intramolecular bonds. **18a:** Density in g/mL helps us understand the mass of a substance per millilitre. **18b:** The amount of substance depends on the mass of substance present/required. **18c:** The energy generated depends on how much of the fuel is used. **19a:** Covalent bonds are a result of atoms sharing electrons. **19b:** Covalent bonds are very strong intramolecular bonds. **19c:** Bromine reacts with unsaturated carbons. **19d:** Diamonds are lattice carbon allotropes. **20a:** Oil is considered a mixture of compounds. **20b:** Not all substances are able to mix. **20c:** Errors can be classified based on the effect it has on the experiment. **20d:** A personal error can be remedied by accounting for the extent of the error.

6C NAMING SIMPLE HYDROCARBONS

In this lesson, we will be introduced to the IUPAC nomenclature of organic chemistry, which provides a systematic process for naming hydrocarbons.

6A Carbon materials	6B Introduction to hydrocarbons	6C Naming simple hydrocarbons	6D Naming complex hydrocarbons	6E Functional groups - alcohols and carboxylic acids	6F Functional groups - esters	6G Polymers	6H Modifying polymers

Study design dot point

- the grouping of hydrocarbon compounds into families (alkanes, alkenes, alkynes, alcohols, carboxylic acids and non-branched esters) based upon similarities in their physical and chemical properties including general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties

Key knowledge units

Naming simple hydrocarbons	1.2.8.3
Naming unsaturated hydrocarbons	1.2.8.5
Naming branched hydrocarbons	1.2.8.4

Key terms and definitions

Systematic name (referring to organic compounds) standardised name

Parent chain longest continuous chain of carbon-carbon bonds. This needs to include any carbon double/triple bonds (ie functional groups) if present

Parent name component of name that refers to the length of the parent chain

Alkyl group a branch associated to the parent carbon chain that contains only carbon and hydrogen atoms

Lesson links

This lesson builds on:

- 6A Carbon materials

Hydrocarbons are a type of carbon compound.

- 6B Introduction to hydrocarbons

Hydrocarbons are named differently depending on their structures.

Naming basic hydrocarbons 1.2.8.3

OVERVIEW

A set of naming conventions determined by IUPAC provide a systematic method of naming any organic molecule.

THEORY DETAILS

In the last lesson we introduced hydrocarbons, exploring the structure, uses and variations of these molecules which are found so abundantly throughout the world. Since there is such a large variety of different hydrocarbons, a standard naming convention is required in order to distinguish between the structures of hydrocarbons through their names alone.

To help ensure that organic molecules are named in a consistent manner, the International Union of Pure and Applied Chemistry (IUPAC) developed a naming system that is the standard for naming organic molecules worldwide. The standardised name for organic compounds prescribed by IUPAC is also referred to as the **systematic name**. There are many rules involved in this naming system, which we will work through in this chapter.

When naming hydrocarbons, we need to first consider the length of the carbon chain, as the longest carbon chain in the molecule is the basis of the name of the molecule. The longest continuous carbon chain in an organic molecule is known as the **parent chain**, and this provides the **parent name** of the molecule. It's important to note that when identifying the parent chain, any functional groups present must be included within the chain. For basic hydrocarbons, the naming convention is shown in figure 1.

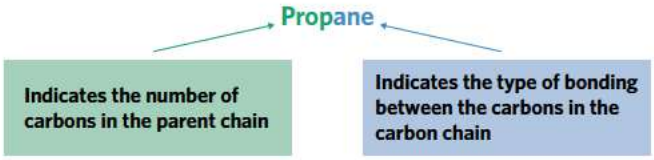


Figure 1 Naming convention of simple hydrocarbons.

As seen above, every component of the name of a hydrocarbon signifies something about the structure of the hydrocarbon itself. As summarised in table 1, a hydrocarbon containing only carbon to carbon single bonds in the parent chain is marked by the suffix **-ane**, whereas a hydrocarbon with one (or more) carbon to carbon double bond in the parent chain is denoted by the suffix **-ene** and a hydrocarbon with one (or more) carbon to carbon triple bonds in the parent chain will contain the suffix **-yne**.

Table 1 Naming conventions of basic hydrocarbons.

Bond type between carbons in the parent chain	Suffix	Type of hydrocarbon	Example
Single	-ane	Alk**ane**	Prop**ane**
Double	-ene	Alk**ene**	Prop**ene**
Triple	-yne	Alk**yne**	Prop**yne**

The parent name of a hydrocarbon is another important component as it identifies the number of carbons in the parent chain. Table 2 summarises the parent names used to indicate the chain length of a hydrocarbon.

Table 2 Hydrocarbon parent names.

Carbon chain length	Parent name	Example
1	Meth-	Methane
2	Eth-	Ethane
3	Prop-	Propane
4	But-	Butane
5	Pent-	Pentane
6	Hex-	Hexane
7	Hept-	Heptane
8	Oct-	Octane
9	Non-	Nonane
10	Dec-	Decane

Through combining the parent name denoting the carbon chain length with the suffix denoting the types of bonds in the molecule, it is possible to come up with the systematic name of many basic hydrocarbons.

Consider the molecule shown in figure 2.

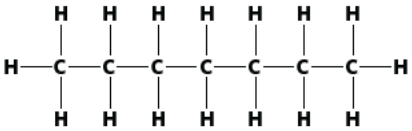


Figure 2 Structural formula for a simple hydrocarbon.

The parent chain can be seen as having a total of seven carbon atoms bonded together in an unbroken chain as shown in figure 3. As such, the parent name for the molecule will be hept-, as can be seen in table 2.

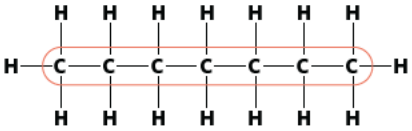
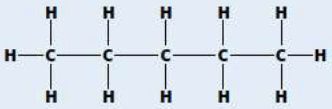


Figure 3 Longest unbroken carbon chain in the hydrocarbon.

Since there are only carbon to carbon single bonds in the parent chain, we classify it as an alkane. Therefore the suffix for this molecule will be -ane. If we combine these components (hept- and -ane) together, the systematic name of this molecule will be heptane. As can be seen, in simple molecules the systematic name can be determined simply by identifying the length of the longest carbon chain and the types of bonds in the molecule. Furthermore, now we can see that the names of these molecules inform the reader on the structure of these molecules, even if they have not seen the molecule before, as the names are produced through a series of standard conventions.

Worked example 1

What is the systematic name of the molecule with the following structural formula?



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

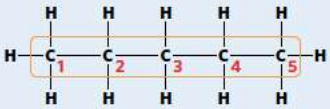
Provide the systematic name of the molecule provided.

What strategy(ies) do we need in order to answer the question?

1. Determine the longest carbon chain length.
2. Determine the types of bonds within the parent carbon chain.
3. Apply this information using correct IUPAC naming conventions.

Answer

This molecule is an alkane, and the longest carbon chain length is five carbon atoms.



The parent name pent- is used to denote five carbon atoms in the parent chain. The molecule itself contains only carbon to carbon single bonds so the suffix -ane is used.

Therefore the name of this molecule is pentane.

It is important to know that hydrocarbons can also have functional groups in their structure. In these cases, the functional groups that the molecule possesses are also denoted through the use of suffixes and prefixes. The naming conventions for these functional groups will be covered more extensively in the next lesson.

Naming unsaturated hydrocarbons 1.2.8.5

OVERVIEW

Hydrocarbons with carbon to carbon double or triple bonds have a more in-depth naming system compared to simple hydrocarbons.

THEORY DETAILS

Unsaturated hydrocarbons have at least one carbon to carbon double or triple bond. The way in which we identify this, as shown in table 1, is to change the suffix of the name accordingly. To be more specific, we also need to indicate where the carbon to carbon double or triple bond is located in the parent chain.

Let's have a look at the molecule shown in figure 4.

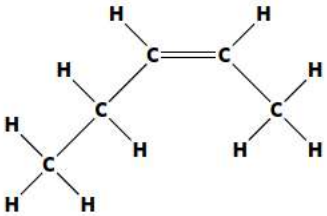


Figure 4 Structural formula of an unsaturated hydrocarbon.

We can see that the longest carbon chain contains five carbons. There is a carbon to carbon double bond in the structure so the suffix for the molecule would be '-ene'. However, we need to be able to indicate to the reader the location of the double bond. To do so, we need to number the carbon atoms in the parent chain. As shown in figure 5, there are two ways in which we can do this depending on which end we start from.

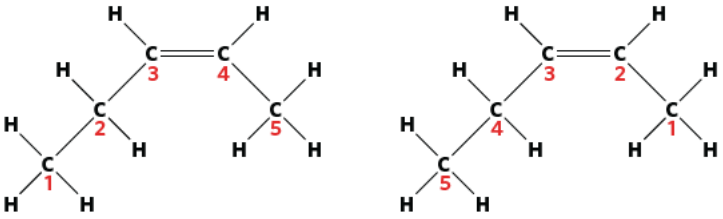


Figure 5 Numbering of the parent chain of a hydrocarbon.

The way in which we decide which numbering system to use is to determine which system allows for the carbon to carbon double bond to occur at the lowest numbered carbon. Based on the two options, we can see that the carbon to carbon double bond can occur at either carbon number 3 or carbon number 2. As such, we will proceed with the numbering system where the carbon to carbon double bond occurs at carbon number 2. The name of this hydrocarbon is shown in figure 6.

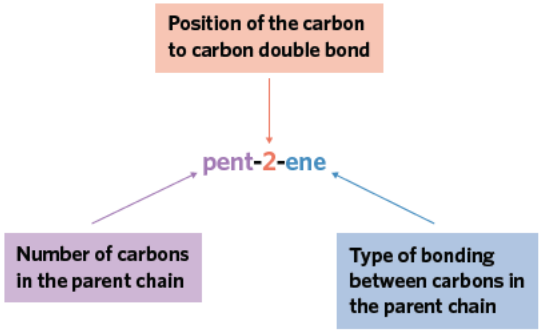


Figure 6 Naming convention of unsaturated hydrocarbon.

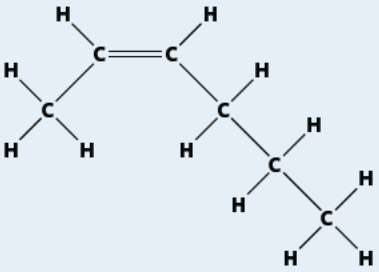
As we can see, the positioning of the carbon to carbon double is indicated just before the suffix. When using numbers to show the position of these bonds, we must use a hyphen between the number and words. This will become important as we progress in our understanding of organic compounds.



Don't forget the hyphen between the numbers and words when writing the name of organic compounds.

Worked example 2

What is the systematic name of the molecule with the following structural formula?



What information is presented in the question?

The structural formula for a molecule.

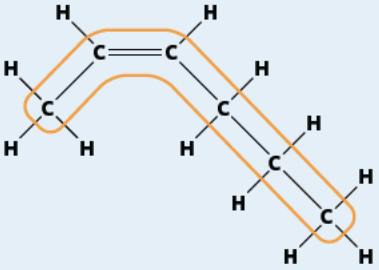
What is the question asking us to do?

Provide the systematic name of the molecule provided.

What strategy(ies) do we need in order to answer the question?

1. Determine the longest carbon chain length.
2. Determine the types of bonds within the parent carbon chain.
3. Number the carbons in the parent chain.
4. Apply this information using correct IUPAC naming conventions.

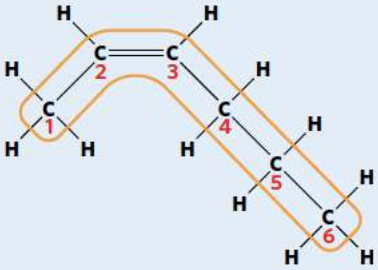
Answer



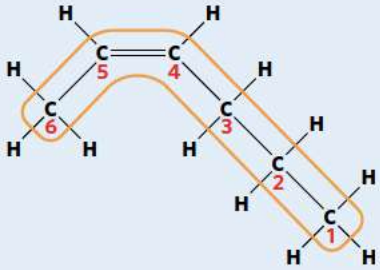
The longest carbon chain contains six carbons. The structure also contains a double bond, and therefore the suffix will be '-ene'.

There are two different ways we can number the carbons in the parent chain:

Option 1



Option 2



Option 1 allows for the carbon to carbon double bond to occur at the lowest number carbon, carbon 2. Therefore the name of this compound is hex-2-ene.

Naming branched hydrocarbons 1.2.8.4

OVERVIEW

Some hydrocarbons have branches associated with the parent chain and therefore require a more in-depth naming system.

THEORY DETAILS

Hydrocarbons can contain side groups/side chains, known as **alkyl groups**, which are composed of carbon and hydrogen atoms, much like the parent chain of the hydrocarbon. They can sometimes also be referred to as branches. Although they have a specialised name, these side groups are also considered as functional groups. These groups can be found bonded to carbons within the parent chain of a hydrocarbon, however as they are not at the end of the parent chain they do not extend the length of the longest continuous carbon chain.



Alkyl side chains are like mini hydrocarbons that are bonded to a main hydrocarbon chain.

The presence of an alkyl group is evident through a prefix added to the name of a hydrocarbon. Similar to the conventions used to name alkanes, alkyl groups have different names depending on the number of carbon atoms in the alkyl group. This is summarised in table 4.

Table 4 Names of common alkyl groups.

Alkyl group	Number of carbons in the group	Parent name	Name of alkyl group
-CH-	1	meth-	methyl
$-CH_2CH_3$	2	eth-	ethyl
$-CH_2CH_2CH_3$	3	prop-	propyl
$-CH_2CH_2CH_2CH_3$	4	but-	butyl
$-CH_2CH_2CH_2CH_2CH_3$	5	pent-	pentyl

When naming hydrocarbon chains with alkyl groups (and really any side groups), we must first identify the number of carbons on the parent chain. Consider the molecule shown in figure 7.

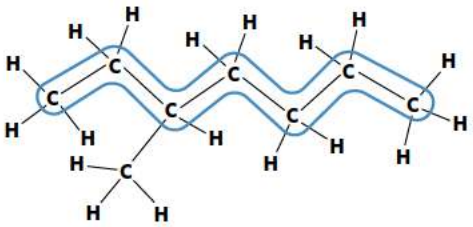


Figure 7 Identification of the parent chain in a hydrocarbon.

As highlighted by the blue circle, the longest carbon chain contains seven carbons. We can also see that the parent chain contains only carbon to carbon single bonds. As such, we can identify that the name of the hydrocarbon is heptane. However, there is another portion of the molecule that we are yet to include in the name of our molecule; the alkyl group found in the orange circle shown in figure 8.

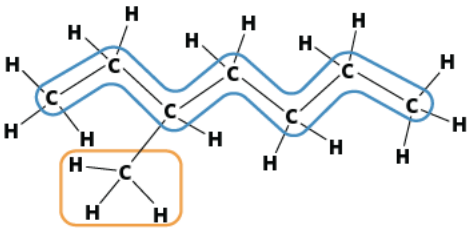


Figure 8 Identification of the alkyl group in a hydrocarbon.

As we can see, the alkyl group contains only one carbon and therefore, as can be seen in table 4, this would be called a methyl group. With this, we can update the name of this compound to methylheptane.

The final piece of information that is missing is where the alkyl group is found. When naming compounds, we need to be able to signal the exact location of side groups (and functional groups) within a molecule. Similar to the way we numbered the carbons in the parent chain when naming unsaturated hydrocarbons, we need to number the carbons in a way that allows the alkyl group to occur at the lowest numbered carbon.

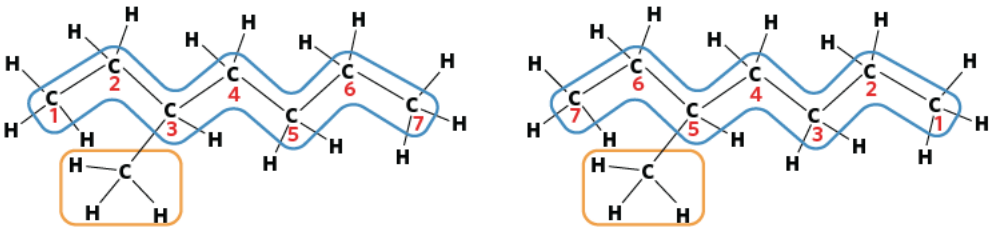


Figure 9 Numbering of the parent chain in a hydrocarbon with a side group.

When comparing the two diagrams in figure 9, the molecule counting from the left has the methyl group bonded to carbon number 3, whereas the molecule counting from the right has the methyl group bonded to carbon number 5. We always want to have our side groups occurring on the lowest numbered carbon possible, therefore in the above example, we would number the parent chain so that the methyl group falls on carbon number 3. When indicating this in the name of the molecule, the name and position of the alkyl group comes before the name of the parent chain.



We number carbon atoms such that side groups are bonded to the lowest numbered carbon possible.

If we combine all of these ideas together, we can identify the name of the compound as shown in figure 10.

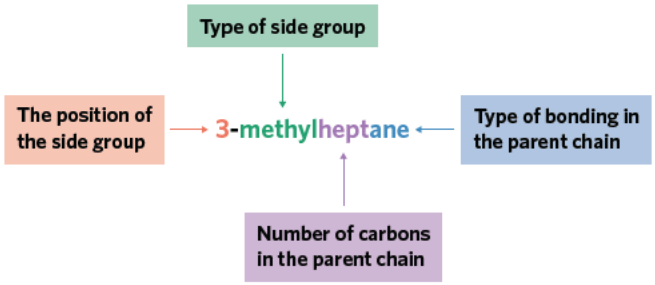
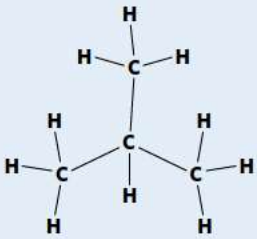


Figure 10 Naming convention of branched hydrocarbon.

As we can see, every part of the name represents a piece of information about the structure of the molecule itself, and therefore it is essential to incorporate every observable component of a molecule into its name to truly identify it.

Worked example 3

What is the systematic name of the following molecule?



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

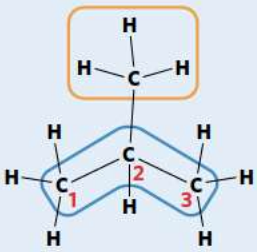
Provide the systematic name of the molecule provided.

What strategy(ies) do we need in order to answer the question?

1. Determine the length of the longest carbon chain.
2. Determine the type of bonding within the parent carbon chain.
3. Identify any side groups/functional groups present.
4. Identify the carbon number to which the side group/functional group is bonded.
5. Name the molecule using IUPAC naming conventions.

Answer

The longest carbon chain contains three carbons, and the only type of bonds within the parent chain is carbon to carbon single bonds. This means that this molecule is an alkane with a parent name of 'prop-'.



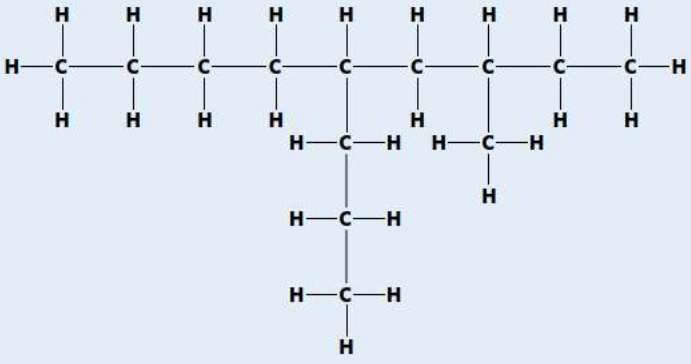
There is a methyl side group found on carbon number 2.

The name of this molecule is 2-methylpropane.

Sometimes, these compounds will have more than one side group bonded to the parent chain. When this occurs, the same rules apply. The only difference is that the alkyl groups are named in alphabetical order, as seen in worked example 4.

Worked example 4

Use the IUPAC naming system to determine the name of the molecule provided.



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

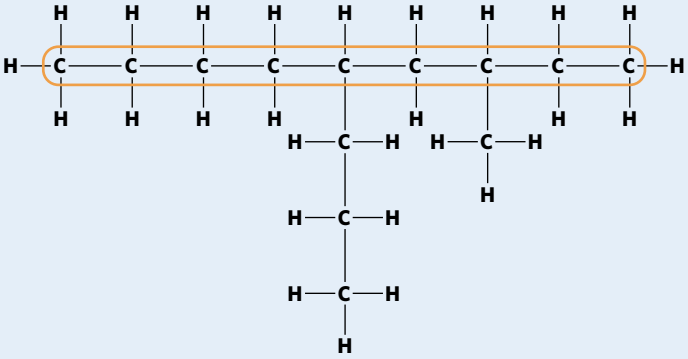
Provide the systematic name of the molecule provided.

What strategy(ies) do we need in order to answer the question?

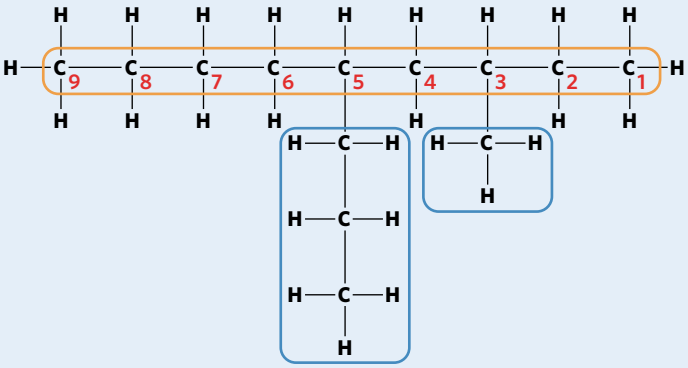
1. Determine the longest carbon chain length.
2. Determine the type of bonding within the parent carbon chain.
3. Identify any side groups/functional groups present.
4. Identify the carbon number to which the side group/functional group is bonded.
5. Name the molecule using IUPAC naming conventions.

Answer

The longest carbon chain contains nine carbons and within the parent chain is only carbon to carbon single bonds. This means that this molecule is an alkane with a parent name of 'non'.



There are two alkyl groups present in the molecule; one with 1 carbon and one with three carbons. As such, this molecule contains methyl and propyl side groups.



Based on the numbering of the carbons in the parent chain, we can see that the methyl group is bonded to carbon number 3 and propyl is bonded to carbon number 5. 'm' comes before 'p' in the alphabet, so methyl is named before propyl.

Therefore, the name of this molecule is 3-methyl-5-propylnonane. This name allows for the side groups to be located at the lowest carbon number.

At times, there may be multiple of the same side groups present on the same hydrocarbon. When this is the case, we use prefixes such as di/tri to indicate multiple of the same groups, with di- indicating two of the side group and tri- indicating three of the side group being present. For example, consider the molecule shown in figure 11.

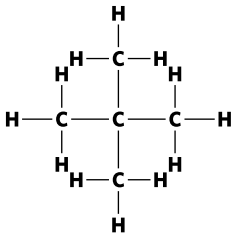


Figure 11 Structural formula of a hydrocarbon with multiple side groups.

Based on our understanding of the naming conventions, we can see that the parent chain contains three carbons. The molecule also contains two of the same side groups located on the same carbon as shown in figure 12.

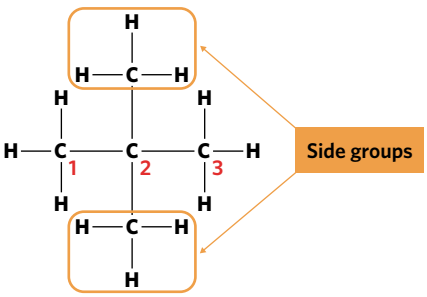


Figure 12 Identification of the side groups found in a branched hydrocarbon.

Since the side groups contain only one carbon, they are both methyl groups. The name of this hydrocarbon is shown in figure 13.

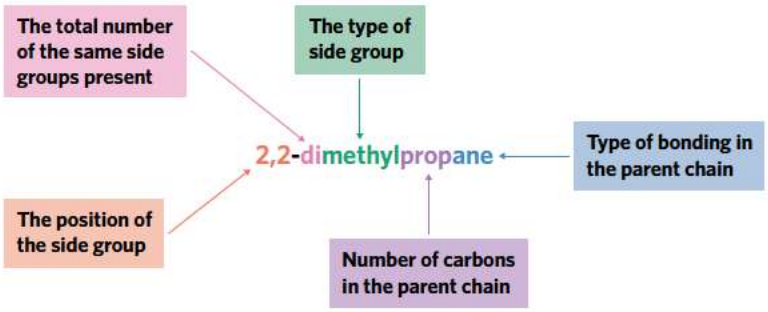


Figure 13 Naming convention of a hydrocarbon containing multiple side groups.



Don't forget to use commas to separate numbers indicating multiple of the same side groups/functional groups.

As we can see, the position of every side group present needs to be indicated in the name of the compound. If multiple of the same side groups or functional groups are present, the location of these are separated by a comma.

Summary of naming simple hydrocarbons

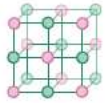
There are many different components of a molecule and observations that must be taken into account when naming hydrocarbons. The following is a summary of all of the steps that should be taken when naming a molecule:

1. Identify the longest unbranched carbon chain, and count the number of carbon atoms to identify the parent name.
2. Number the positions of the carbons in the parent chain. This will be affected by the priority of the side groups present.
3. Identify the number and position(s) of the side group.
4. Identify the position(s) and name(s) of the side group, using the prefix 'di-/tri-' in the case that it is present multiple times.
5. Combine the above to write the complete name of the organic compound.

As can be seen, the name of the hydrocarbon is fully dependent on the length of the carbon chain and substituents present in the molecule. In the next lesson we will learn how to put all of these elements together to name more complex hydrocarbons.

Theory summary

- The systematic name of a molecule is developed based on the IUPAC naming system.
- The parent chain represents the longest continuous chain of carbon atoms bonded together.
- The parent name of a hydrocarbon signifies the number of carbons in the parent chain.
- The suffix of the name of a hydrocarbon can indicate whether the hydrocarbon is an alkane, alkene, or alkyne.
- Alkyl groups are branches bonded to carbons in the parent chain.
- Alkyl groups consist of hydrogen and carbon atoms.
- Alkyl groups are named based on the number of carbons.



6C QUESTIONS

Theory review questions

Question 1

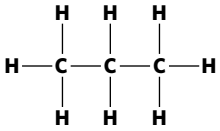
The suffix used to indicate a double bond is

A -ene.

B -yne.

Question 2

The parent name of this molecule is

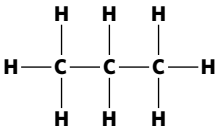


A prop-.

B but-.

Question 3

The suffix for the hydrocarbon shown is

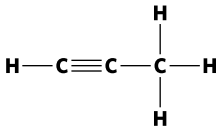


A -ene.

B -ane.

Question 4

The suffix for the hydrocarbon shown is



A -ene.

B -yne.

Question 5

The name of the molecule shown in question 4 is

A prop-2-yne.

B prop-1-yne.

Question 6

The alkyl group involving two carbons atoms is called

A methyl.

B ethyl.

Question 7

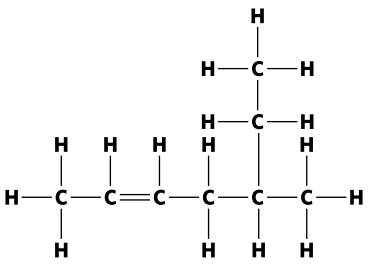
Which of the following molecules contains more carbon atoms?

A 3-methyloctane

B 3,3-diethylhexane

Question 8

The name of the side group in this molecule is

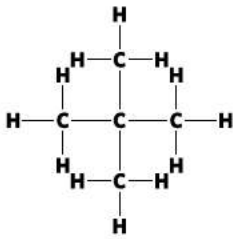


A ethyl.

B methyl.

Question 9

To denote the two methyl groups found in the molecule shown, we would use



A methyl-methyl – at the start of the hydrocarbon name.

B dimethyl – at the start of the hydrocarbon name.

Question 10

Since the methyl groups of the molecule in question 9 are found on the same carbon, we would denote this as

A 2-dimethyl.

B 2,2-dimethyl.

Deconstructed questions

Use the following information to answer questions 11–13.

Dorian is a gentleman with a peculiar interest in oil paints, and wishes to understand the hydrocarbons found in them. In order to achieve this, he must understand the naming conventions of hydrocarbons, as designed by The International Union of Pure and Applied Chemistry such that people can understand the structure of a hydrocarbon through features of its name alone.

Question 11 (1 MARK)

Which of the following options has the parent names all associated with the correct number of carbon atoms in the carbon chain?

A Pent-: 5, Oct-: 8, Prop-: 4

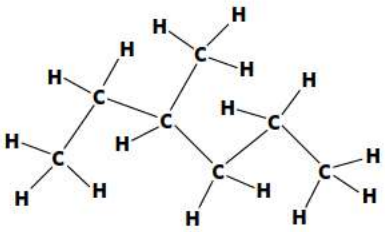
B Pent-: 3, Oct-: 8, Prop-: 5

C Pent-: 3, Oct-: 7, Prop-: 5

D Pent-: 5,: Oct-: 8, Prop-: 3

Question 12 (1 MARK)

Which of the following is the correct IUPAC name of the molecule shown?



A 2-ethylpent-4-ene

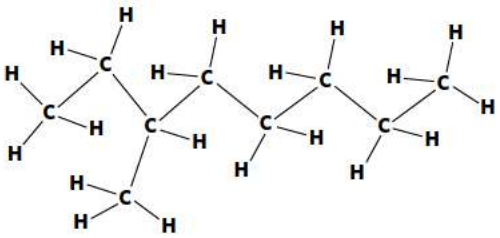
B 4-ethylpent-1-ene

C 4-methylhexane

D 3-methylhexane

Question 13 (3 MARKS)

Having learnt about the naming of carbon chain lengths and functional groups in hydrocarbons, Dorian names the molecule shown as 2-ethylheptane:



Explain why Dorian's suggested name for this molecule is incorrect, and provide the correct name.

Exam-style questions

Within lesson

Question 14 (6 MARKS)

Professor Burkley, a renowned geologist, has been prompted to answer a series of challenges regarding the carbon chain structure of linear hydrocarbons.

a Give the names of the single-chain saturated hydrocarbons containing:

i 5 carbon atoms (1 MARK)

ii 8 carbon atoms (1 MARK)

iii 4 carbon atoms (1 MARK)

b Write the following molecules in order of decreasing carbon chain length: Octane; decane; heptane; propane; ethane; hexane (1 MARK)

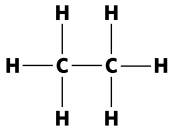
c What is the total number of carbon atoms in the molecule 3-methylhexane? Explain your answer. (2 MARKS)

Question 15 (7 MARKS)

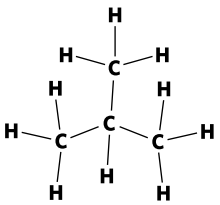
A chemistry professor has made the error of losing the labels for a number of different containers with hydrocarbons. Knowing only the molecular formulae, he is currently relabelling the hydrocarbons.

a Provide the IUPAC systematic names for each of the following molecules:

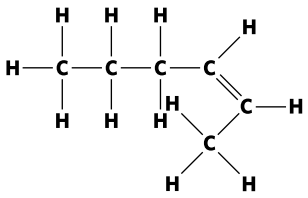
i (1 MARK)



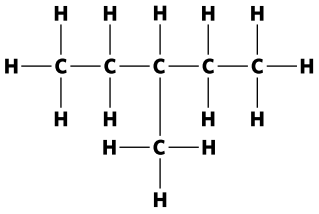
ii (1 MARK)



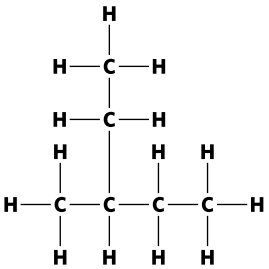
iii (1 MARK)



iv (1 MARK)



b The professor labelled the following molecule as 2-ethylbutane.



i Explain why the name given by the professor is incorrect. (2 MARKS)

ii Provide the IUPAC systematic name of this molecule. (1 MARK)

Question 16 (8 MARKS)

Lucinda is a year 11 chemistry student learning about the structures of hydrocarbons. Having learned the systematic IUPAC naming conventions of hydrocarbons, she now needs to connect this information to the ability to draw structural formulas of hydrocarbons.

a Draw the structural formulas of the following hydrocarbons.

i But-2-ene (1 MARK)

ii 3-methylpentane (1 MARK)

iii 2,3-dimethylheptane (1 MARK)

While learning about the naming conventions of hydrocarbons, Lucinda came up with the molecule '6-methyloctane', however she has her doubts about this molecule.

b Explain why the molecule 6-methyloctane could not exist. (2 MARKS)

c What is the correct name of the molecule Lucinda was attempting to name? (1 MARK)

d Draw the structural formula of the molecule named in 16c. (2 MARKS)

Multiple lessons

Question 17 (4 MARKS)

The molecular formula of a hydrocarbon is C_6H_{14}.

a What is meant by the term 'isomer'. (1 MARK)

b Explain why the structural isomer 1,1-dimethylbutane could not exist, and provide the correct IUPAC name of this molecule. (2 MARKS)

c Identify the name of the molecule with this molecular formula and no side chains. (1 MARK)

Question 18 (8 MARKS)

Hydrocarbons are organic molecules found in abundance worldwide which require standard naming conventions.

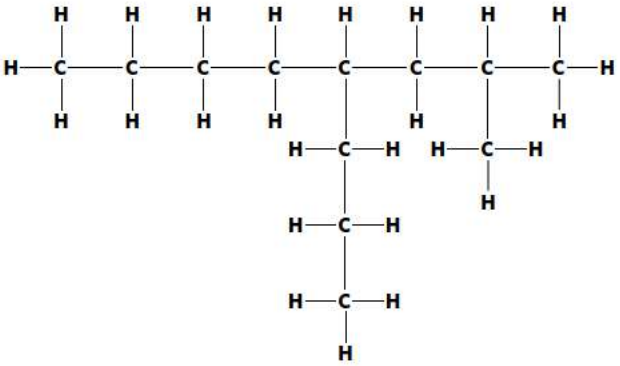
a What is meant by the term 'organic molecule'. (1 MARK)

b Hydrocarbons are members of a homologous series of molecules containing carbon and hydrogen.

i Using at least 2 diagrams, explain the term 'homologous series' and how it relates to hydrocarbons. (3 MARKS)

ii When naming hydrocarbons of the same homologous series, what do they have in common? (1 MARK)

c What is the IUPAC systematic name of the molecule shown? Explain how you produced this answer. (3 MARKS)



Question 19 (10 MARKS)

Alandra is a chemist tasked with the job of labelling a series of unknown substances. The method used in this procedure is to determine the molecular formula of the substance through chemical analysis, and to determine the IUPAC systematic name of the molecule from this.

a Categorise the following molecules as alkanes, alkenes or alkynes.

i C_4H_{10} (1 MARK)

ii C_7H_{16} (1 MARK)

iii C_3H_4 (1 MARK)

iv C_5H_{10} (1 MARK)

b It is known that some of the molecules contain alkyl groups, and therefore structural isomers must be accounted for. Draw and name all the structural isomers of the molecules with the following molecular formula C_5H_{12}. (6 MARKS)

Key science skills

Question 20 (6 MARKS)

Theo is performing chemical analysis on a sample of various hydrocarbons – heptane, hexane and nonane. As part of his analysis, he wanted to assess the properties of each, including boiling point, viscosity and heat of combustion.

a When calculating the boiling points of each of these fuels, Theo's results were not what he expected. Although all of the results were close to each other, none of them were close to the boiling point he expected.

i Comment on the accuracy and precision of the results. (2 MARKS)

ii What type of error do these results indicate? Explain your reasoning. (2 MARKS)

b Calculating the boiling point of a substance involves the use of an open flame. Suggest two safety precautions that should be taken to prevent injury. (2 MARKS)

Exam-style question hints

14a: Prefixes are used to signify the number of carbons in the parent chain. **14b:** Prefixes are used to signify the number of carbons in the parent chain. **14c:** Side groups can contribute carbon atoms to the overall structure of the molecule. **15a:** The naming of hydrocarbons is heavily dependent on the bond type and hydrocarbon chain. **15bi:** Naming of a molecule involves determining the longest carbon chain. **15bii:** The name of the alkyl group is written first. **16a:** Each section of a hydrocarbon's name identifies a key structural feature of the molecule. **16b:** The position of a branch is determined based on the shortest distance from the end/main carbon. **16c:** The position of a branch is determined based on the shortest distance from the end/main carbon. **16d:** Structural formulas show all covalent bonds **17a:** Isomers share a common characteristic **17b:** The parent chain of a molecule is the longest continuous carbon chain. **17c:** Each type of hydrocarbon has a different ratio of carbon to hydrogen. **18a:** Organic does not mean natural. **18bi:** Hydrocarbons are based around their carbon chains. **18bii:** The name of hydrocarbons are based on their structure. **18c:** The IUPAC naming system uses the structure of the molecule as the point of reference. **19a:** Each type of hydrocarbon follows a general formula. **19b:** Structural isomers have the same atoms in different spatial orientations. **20ai:** Accuracy and precision describe different qualities of a test's results. **20aii:** Precise results are all similar to each other. **20b:** An open flame can cause safety concerns.

6D NAMING COMPLEX HYDROCARBONS

In this lesson, we will apply our knowledge of the IUPAC naming system to more complex hydrocarbons.

6A Carbon materials	6B Introduction to hydrocarbons	6C Naming simple hydrocarbons	6D Naming complex hydrocarbons	6E Functional groups - alcohols and carboxylic acids	6F Functional groups - esters	6G Polymers	6H Modifying polymers

Study design dot point

- the grouping of hydrocarbon compounds into families (alkanes, alkenes, alkynes, alcohols, carboxylic acids and non-branched esters) based upon similarities in their physical and chemical properties including general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties

Key knowledge units

Naming complex hydrocarbons	1.2.8.6
Formulas of hydrocarbons	1.2.8.7

Naming complex hydrocarbons 1.2.8.6

OVERVIEW

Some hydrocarbons have multiple different components that need to be indicated the name of the compound.

THEORY DETAILS

Some hydrocarbons can be unsaturated as well as containing branches.

Consider the compound in figure 1. The parent chain contains five carbon atoms and one carbon to carbon double bond. As shown in figure 2, we can also see that there are two methyl groups associated with the parent chain.

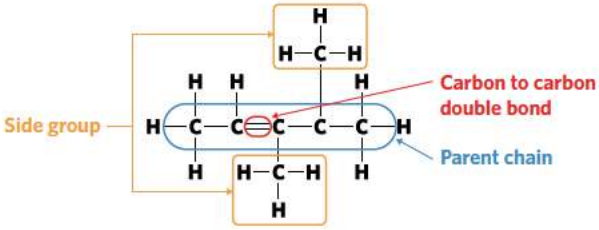


Figure 2 Breakdown of a structural formula.

In previous examples, we have numbered the carbons in the carbon chain based on the location where the side groups can be found. However, in this example, there is a carbon to carbon double bond in addition to the side groups. As we know, this is also considered a functional group. When naming molecules, functional groups take priority and therefore, in this instance we number the carbons in the parent chain in a way that the carbon to carbon double bond occurs the earliest.



Figure 3 Numbering of the carbons in a parent chain.

If we just consider the parent chain, we can number the carbons from both directions as shown in figure 3. One naming system allows the carbon to carbon double bond to occur at carbon 2 whilst the other at carbon 3. As such, we will choose the numbering system where the carbon to carbon double bond occurs at carbon 2 as shown in figure 4.

Lesson links

This lesson builds on:

- 6B Introduction to hydrocarbons

Hydrocarbons are named differently depending on their structures.

- 6C Naming simple hydrocarbons

Complex hydrocarbons can contain carbon to carbon multiple bonds and alkyl groups.

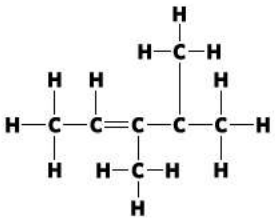


Figure 1 Structural formula of a compound.

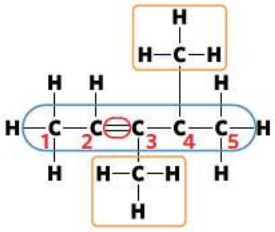


Figure 4 Breakdown of a structural formula.

As we can see, there are also two methyl groups present in the molecule, which fall on carbon numbers 3 and 4. As a result, the name of this hydrocarbon would be:

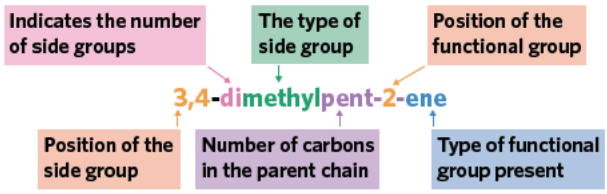
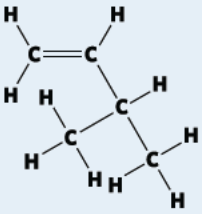


Figure 5 Breakdown of the name of organic compounds.

As we can see in figure 5, when we introduce functional groups and side groups to hydrocarbons, naming becomes more difficult than naming simple alkane hydrocarbon chains. Nonetheless, IUPAC naming conventions can still be used whenever we need to name any type of organic compound.

Worked example 1

Identify the IUPAC name for the organic compound.



What information is presented in the question?

The structure of a hydrocarbon.

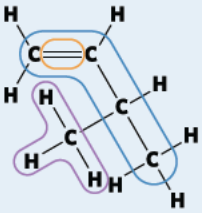
What is the question asking us to do?

Name the hydrocarbon.

What strategy(ies) do we need in order to answer the question?

1. Identify the number of carbons in the parent chain.
2. Identify the type(s) of bonds in the parent chain.
3. Identify any branches.
4. Number the carbons in the parent chain.
5. Apply this information using correct IUPAC naming conventions.

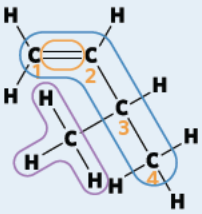
Answer



As shown by the blue line, the parent chain contains four carbons, hence will have a prefix of 'but'.

Shown in the orange line is a carbon to carbon double bond, making this compound an alkene. Therefore the suffix for this compound is '-ene'.

Shown in the purple is a branch/alkyl group containing one carbon. Therefore the alkyl's name is methyl.



Based on the numbering system that allows the carbon to carbon double bond to occur at the lowest numbered carbon, the carbon to carbon double bond occurs at carbon 1 and the methyl group occurs at carbon 3.

Therefore the name of this organic compound is 3-methylbut-1-ene.

Formulas of hydrocarbons 1.2.8.7

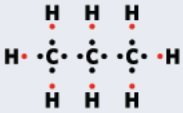
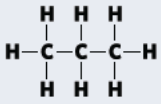
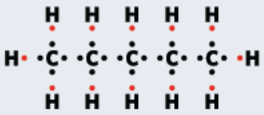
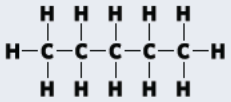
OVERVIEW

Hydrocarbons can be presented using different formulas.

THEORY REVIEW

There are a variety of ways to write hydrocarbon compound formulas. These are shown in table 1.

Table 1 Different representations of hydrocarbons.

Hydrocarbon	Molecular formula	Semi-structural formula	Electron dot formula	Structural formula	Skeletal formula
Propane	C_3H_8	$CH_3CH_2CH_3$			
Pentane	C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$			

We have come across most of these formulas in previous lessons, particularly structural formulas. Although semi structural formulas seem long, they are very useful in helping to identify the structural formula of the compound as shown in figure 6.

As we can see, the order in which the semi structural formula is written reflects the way in which the structural formula would be drawn. Though it doesn't include the single bonds in the structure, it does include any double (and even triple) bonds present. If the structure contains a branch, the alkyl group is generally written brackets. Considering the molecule shown in figure 7. The semi- structural formula would be written as $CH_3CH(CH_3)CH_2CH_3$, where the methyl group is written in brackets.

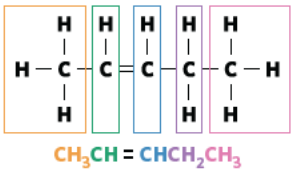


Figure 6 Relationship between structural and semi-structural formulas.

Most of the molecules are represented as molecular formulas, semi structural formulas and structural formulas however VCAA does require that we have a brief understanding of a skeletal formula. In a skeletal structure, as seen in figure 8, the carbons and hydrogens are removed from the structure, leaving only the backbone of the organic molecule. We must then imagine the carbon and hydrogen atoms in the structure.

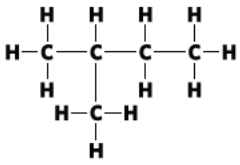


Figure 7 Structural formula for 2-methylbutane.

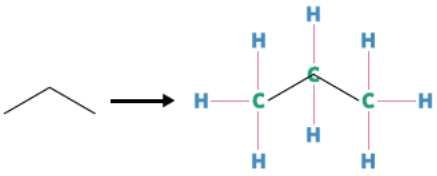
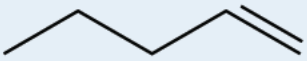


Figure 8 The relationship between skeletal and structural formulas.

Worked example 2

What is the systematic name of the molecule with the following skeletal formula?



What information is presented in the question?

The skeletal formula for a molecule.

What is the question asking us to do?

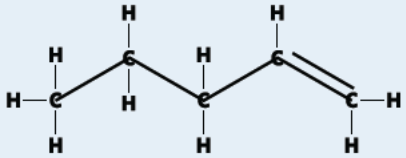
Provide the systematic name of the molecule provided.

What strategy(ies) do we need in order to answer the question?

1. Determine the longest carbon chain length.
2. Determine the type of bonding within the parent carbon chain.
3. Identify any side groups/functional groups present.
4. Identify the carbon number to which the side group/functional group is found.
5. Name the molecule.

Answer

If we imagine the full structural formula of this molecule, it would look like this:



The parent chain contains five carbons. There is also a carbon to carbon double bond located at carbon number 1.

The systematic name of this molecule is pent-1-ene.

Summary of naming hydrocarbons

There are many different aspects of a molecule and observations that must be taken into account when naming hydrocarbons. The following is a summary of all of the steps that should be taken when naming a complex hydrocarbon:

1 Identify the longest unbranched carbon chain, and count the number of carbon atoms to identify the parent name. If the parent chain contains carbon to carbon double or triple bonds (functional groups), this must be included in the parent chain.
2 Identify the type of bonding between the carbon atoms within the parent chain to determine the suffix of the hydrocarbon.
3 Identify the side group(s) present (if any) and thereby the prefix (start) of the name.
4 Number the positions of the carbons in the parent chain. This will be affected by the priority of any carbon to carbon double or triple bonds (functional groups) and side groups present.
5 Identify the number and position(s) of any present functional groups.
6 Identify the position(s) and name(s) of the side group(s), using the prefix 'di-/tri-' in the case that it is present multiple times.
7 Combine the above to write the complete name of the organic compound. Ensure that commas are used to separate the position of multiple of the same alkyl groups and that hyphens are used to separate words and numbers.

As can be seen, the name of the hydrocarbon is fully dependent on the length of the carbon chain and substituents present in the molecule. In the next lesson we will learn about specific functional groups in detail. Regardless of the functional groups present, it is important to understand that hydrocarbons always follow the IUPAC systematic naming conventions, simply adding additional components to the names to account for additional components of a molecule.

Theory summary

- The systematic name of a molecule is developed based on the IUPAC naming system
- The parent chain represents the longest continuous chain of carbon atoms bonded together.
- When numbering the carbons in the parent chain, we number the carbons in a way that prioritises any carbon to carbon double/triple bonds (functional group) over any side groups.

6D QUESTIONS

Theory review questions

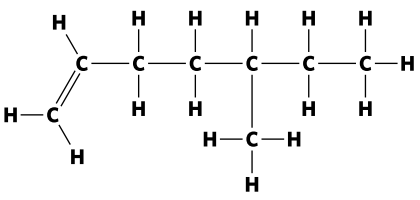
Question 1

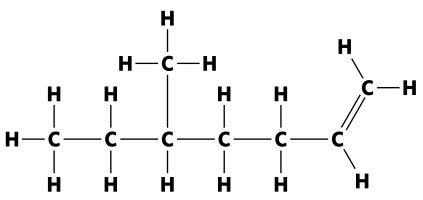
When numbering carbons in an unsaturated hydrocarbon

A alkyl groups are prioritised.
B carbon to carbon double/triple bonds are prioritised.

Question 2

The molecules shown have

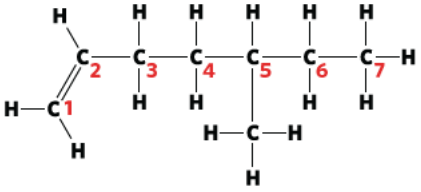




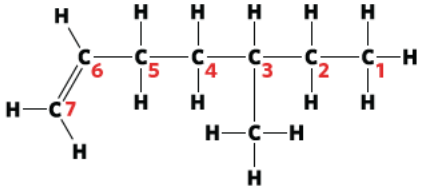
A the same names.
B different names.

Question 3

Which of the following options shows the correct numbering of carbons in the parent chain?



Option 1



Option 2

A Option 1

B Option 2

Question 4

The systematic name of the molecule shown in question 3 is

A 3-methylhept-6-ene.

B 5-methylhept-1-ene.

Question 5

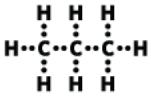
For an unsaturated hydrocarbon containing a methyl group, which of the following is most likely the name of the organic compound?

A 4-methylpent-2-ene

B 2-methylpent-4-ene

Question 6

What is the name of the molecular representation shown?



A Skeletal formula

B Electron dot formula

Question 7

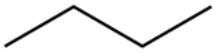
Which of the following structures represents the skeletal structure for the formula shown in question 6?

A $CH_3—CH_2—CH_3$

B 

Question 8

This molecule has



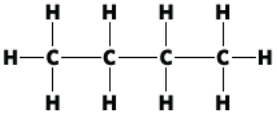
A four carbons.

B no hydrogens.

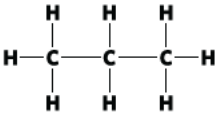
Question 9

Which of the following represents the structural formula for the molecule shown in question 8?

A

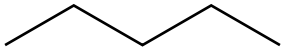


B



Question 10

The condensed formula for the molecule shown is



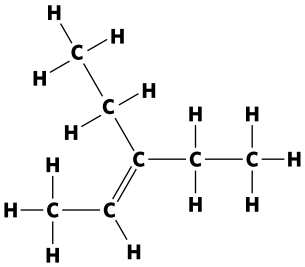
A $CH_3CH_2CH_2CH_2CH_3$.

B $CH_3CH_2CH_2CH_3$.

Deconstructed questions

Use the following information to answer questions 11–13.

Organic compounds have a variety of different uses. Consider the following organic molecule.



Question 11 (1 MARK)

The number of carbons in the parent chain is:

A 4

B 5

C 6

D 7

Question 12 (1 MARK)

The parent name of the molecule would be:

A But-

B Pent-

C Hex-

D Hept-

Question 13 (1 MARK)

Give the systematic name for the molecule provided.

Exam-style questions

Within lesson

Question 14 (1 MARK)

The organic compound with the semi-structural formula is $CH_3CH(CH_3)CH_2CH_3$ will have the systematic name:

A Pentane

B 2-Methylpentane

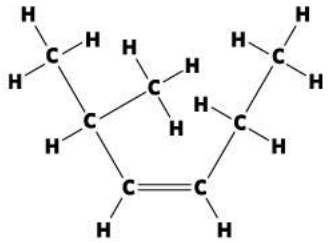
C Butane

D 2-Methylbutane

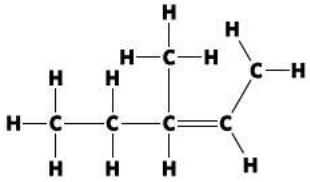
Question 15 (1 MARK)

As a structural formula, 3-methylpent-2-ene would look like:

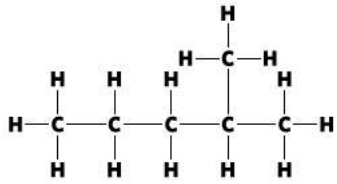
A



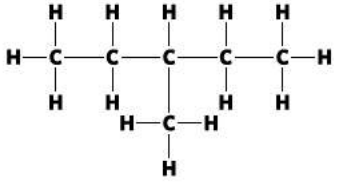
B



C



D

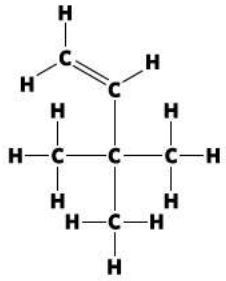


Question 16 (4 MARKS)

Give the systematic name for each of the following compounds:

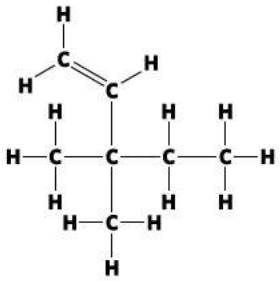
a $CH_3CH_2CH = C(CH_3)CH_3$ (1 MARK)

b



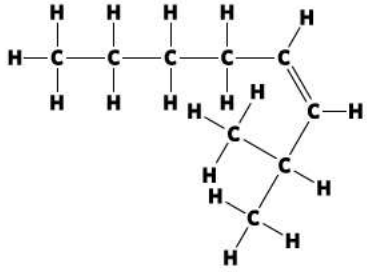
(1 MARK)

c



(1 MARK)

d



(1 MARK)

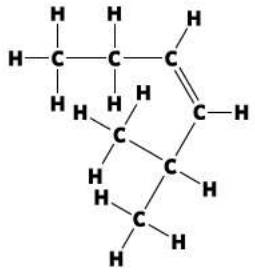
Multiple lessons

Question 17 (4 MARKS)

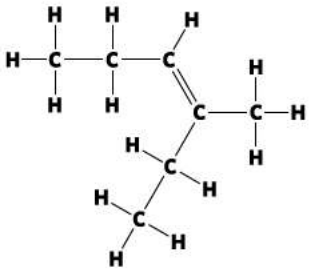
One of the reasons why we are able to have so many different organic compounds is due to the fact that molecules can have the same formula, but different structures.

a What is the term used to describe this attribute? (1 MARK)

b Both of the molecules presented have the same molecular formula.



Molecule A



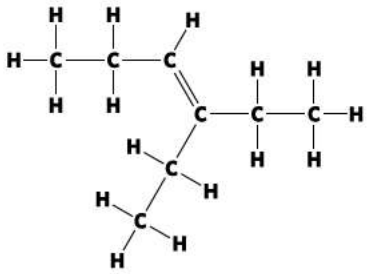
Molecule B

i What is the molecular formula for both molecules? (1 MARK)

ii Identify the systematic name for both molecules. (2 MARKS)

Question 18 (4 MARKS)

Compared to other molecules with the same molecular formula, this molecule has a relatively low melting point.



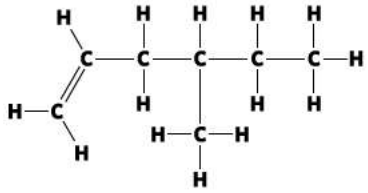
a Determine the molecular mass for the molecule shown. (1 MARK)

b Give the systematic name of the molecule shown. (1 MARK)

c Would this molecule be considered a saturated or unsaturated hydrocarbon? Explain. (2 MARKS)

Question 19 (5 MARKS)

Alkenes are able to take part in addition reactions with hydrogen gas to form alkanes.



a Based on the structure of the molecule, what type(s) of intermolecular bonds would need to be broken in order to change the substance from a liquid to gaseous form? Explain (3 MARKS).

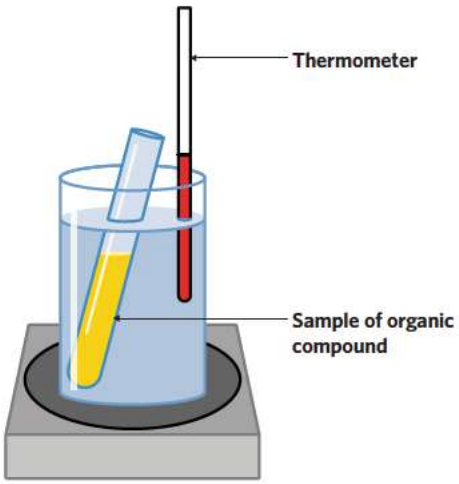
b Give the semi-structural formula for the molecule shown. (1 MARK)

c In a reaction with hydrogen, a total of 0.122 mol of the compound was used during the reaction. Calculate the mass, in grams, of the compound used. (1 MARK)

Key science skills

Question 20 (7 MARKS)

Different organic compounds have different boiling points depending on their structure. Some students wanted to test the boiling points of three different organic compounds: methylhexene, propane and hept-2-ene. The students set up the experiment using the following method:



1. Add 300mL of water into a 500mL beaker.
2. Place the beaker on top of an electric hot plate.
3. Add a 20mL sample of an organic compound into a test tube.
4. Place the test tube in the beaker.
5. Turn on the electric hot plate and allow the beaker of water to heat.
6. Observe until the substance in the test tube starts to boil.
7. Use a thermometer to measure the temperature of the water when the substance in the test tube starts to boil.
8. Record results.
9. Repeat steps 3–8 for the remaining samples.

a Identify the dependent variable. (1 MARK)

b Of the organic molecules tested, which would be considered a saturated hydrocarbon? Explain your answer. (2 MARKS)

c The students realised that they had misread the experimental method and in their experiment allowed for the test tube to touch the bottom of the beaker. This resulted in the test tube being within close proximity to the heating pad of the hot plate. Explain the impact that this would have on boiling points measured. (4 MARKS)

Exam-style question hints

14: The bracket in a semi-structural formula signifies side groups bonded to the parent chain. **15:** The IUPAC naming system will always present carbon to carbon double/triple bonds and side chains at the lowest numbered carbon. **16:** Carbon to carbon double bonds must be included in the parent chain. **17a:** A special name is given to compounds with the same formula but different structural formulas. **17bi:** The molecular formula identifies all the atoms of each element present in a molecule. **17bii:** The IUPAC naming system will always present carbon to carbon double/triple bonds and side chains at the lowest numbered carbon. **18a:** The molecular mass is the mass of a molecule of the substance **18b:** Like carbon to carbon double bonds, the position of alkyl groups is on the lowest numbered carbon possible. **18c:** Saturation of hydrocarbons depends on the type of bonding in the parent chain. **19a:** Intermolecular bonds depend on the type of covalent bonding present within the molecule. **19b:** Semi-structural formulas do not include single bonds. **19c:** The mass of the compound present depends on the amount of substance available. **20a:** The dependent variable depends on the variable that is changed in the experiment. **20b:** Saturation depends on the type of bonding within the parent chain. **20c:** In a beaker of water, the thermometer measures the average temperature of the liquid.

6E FUNCTIONAL GROUPS - ALCOHOLS AND CARBOXYLIC ACIDS

In this lesson, we will be looking at the functional groups of hydrocarbon compounds in greater detail. Particularly, we will be investigating functional groups associated with alcohols and carboxylic acids.

6A Carbon materials	6B Introduction to hydrocarbons	6C Naming simple hydrocarbons	6D Naming complex hydrocarbons	6E Functional groups - alcohols and carboxylic acids	6F Functional groups - esters	6G Polymers	6H Modifying polymers

Study design dot point

- the grouping of hydrocarbon compounds into families (alkanes, alkenes, alkynes, alcohols, carboxylic acids and non-branched esters) based upon similarities in their physical and chemical properties including general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties

Key knowledge units

Alcohols	1.2.8.8
Naming alcohols	1.2.8.9
Carboxylic acids	1.2.8.10
Naming carboxylic acids	1.2.8.11

Key terms and definitions

Functional group specific groups of atoms within a compound that affect the properties of the compound

Hydroxyl group functional group consisting of an oxygen atom covalently bonded to a hydrogen atom (-OH)

Alcohol organic compound that has a hydroxyl (-OH) group bonded to a carbon atom in the carbon chain

Structural isomers isomers that have the same molecular formula but the atoms are arranged in different orders

Carboxylic acid organic compound that contains a carboxyl functional group (-COOH)

Carboxyl functional group functional group that has a carbon double bonded to an oxygen and also singly bonded to a hydroxyl group

Lesson links

This lesson builds on:

- 5C Intermolecular bonding

Functional groups alter the intermolecular interactions of molecules.

- 6C Naming simple hydrocarbons

Functional groups change the naming conventions of hydrocarbons.

Alcohols 1.2.8.8

OVERVIEW

Alcohols are organic compounds that possess a hydroxyl functional group.

THEORY DETAILS

In the last lesson, we discussed the basic structure of hydrocarbons, including naming conventions as well as variations in bonding. These hydrocarbons can further diversify with the addition of **functional groups**. These functional groups are defined as specific groups of atoms bound to a hydrocarbon that provide it with some distinct physical and chemical properties. In this lesson, we will be discussing various functional groups, starting with the **hydroxyl group**, which produces **alcohols**.

Alcohols are organic compounds that contain a hydroxyl functional group. This hydroxyl functional group is made up of an oxygen atom covalently bonded to a hydrogen atom, denoted as -OH, and this replaces a hydrogen atom on the carbon chain. An example of an alcohol is shown in figure 1, as well as different representations of alcohols in table 1.

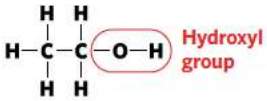
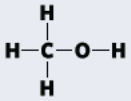
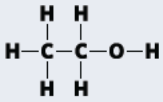


Figure 1 Structure of ethanol, an alcohol.

Table 1 Representations of alcohols.

Name	Molecular formula	Structural formula	Semi-structural formula	Lewis structure
Methanol	CH_3OH		CH_3OH	
Ethanol	C_2H_5OH		CH_3CH_2OH	



The presence of a hydroxyl group signals that the molecule is an alcohol. The functional group itself is NOT referred to as an alcohol.

Properties and uses of alcohols

The presence of a hydroxyl group in an alcohol molecule results in the boiling point of alcohols being higher than that of hydrocarbons that don't contain a hydroxyl group. Due to the difference in electronegativity between oxygen and hydrogen, the bond that occurs between the two atoms in the hydroxyl group is considered a polar covalent bond. Consequently, a hydroxyl group of an alcohol is able to take part in permanent dipole-dipole interactions as well as hydrogen bonding with a hydroxyl group from another alcohol. Such hydrogen bonding is shown in figure 2.

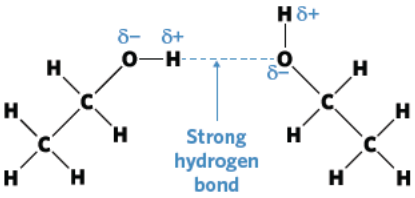


Figure 2 Hydrogen bonding between two ethanol molecules.

Due to the fact that dipole-dipole interactions and hydrogen bonds are relatively strong intermolecular bonds, the amount of energy needed to break these bonds is higher than its corresponding alkane, which is only held together by weaker dispersion forces. Since the intermolecular forces between alcohols are stronger, the boiling points of alcohols are higher than their corresponding alkanes as more energy is required to overcome the intermolecular forces between alcohols than the corresponding alkane. This relationship can be seen in figure 3.

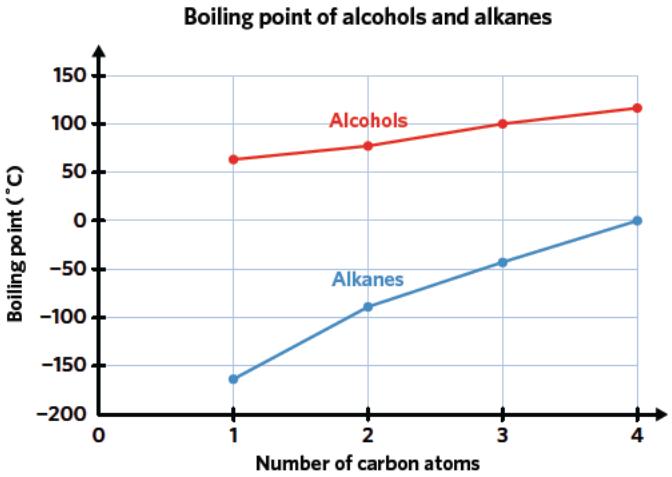


Figure 3 Comparison of the boiling point of alkanes and alcohol.

As we learned previously, the boiling point of alkanes increases as the length of the hydrocarbon's carbon chain increases, due to the increase in dispersion forces. Therefore, while alcohols are able to take part in stronger intermolecular forces than their counterpart alkanes, it is important to note that as the non-polar section of the hydrocarbon chain increases, the effect of the hydroxyl group on the boiling point slowly decreases. We can see in figure 3 that as the number of carbon atoms in the molecule increases, the difference in boiling point between the alcohols and the alkanes decreases as the number of carbon atoms in the chain increases. This shows the decreasing influence of the hydroxyl functional group on the boiling point as the size of the non-polar section grows.

The diminishing influence of the functional group can be pictured as a billboard on top of a building: if the building is only one storey, then the billboard will be very noticeable, however, on top of a skyscraper, the billboard will be far less easy to see. This analogy is shown in figure 4.

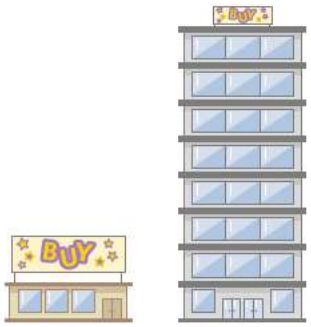


Figure 4 Decreasing influence of functional groups as carbon chain length increases.

While the presence of a hydroxyl group has a significant influence on the intermolecular interactions between alcohol molecules, this also impacts the interactions of alcohols with other types of molecules. The hydroxyl group enables alcohols to form dipole-dipole interactions with other polar molecules, as well as strong hydrogen bonds with water molecules as shown in figure 5, enabling alcohols to dissolve in water.

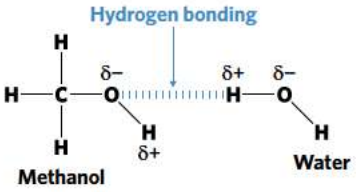


Figure 5 Intermolecular bonds formed between water molecules and methanol.

By comparison, alkanes are non-polar molecules and therefore are unable to form dipole-dipole interactions with water, which means that they cannot dissolve in water. The mechanisms involved in the process of molecules dissolving will be covered in greater detail in a later lesson.

Similarly to the diminishing influence of the hydroxyl group on boiling point, as the carbon chain length of an alcohol increases, the solubility of the molecule in water also decreases. This is because the molecule becomes less polar as the non-polar carbon chain length increases, therefore reducing its ability to form strong dipole-dipole interactions and hydrogen bonds with water molecules.

There are many different alcohols that exist, all of which have different uses, however the primary use of alcohols in modern society tends to be as a fuel. Bioethanol, a form of ethanol produced from the fermentation of glucose, is often added to petrol due to it being renewable. This concept will be covered in greater detail in Unit 3 of VCE Chemistry.

Structural isomers of alcohols

As the definition of an alcohol is a hydrocarbon with a hydroxyl functional group attached, the location of this functional group on the molecule can vary, even though the alcohols have the same formula. Some examples of this are shown in figure 6.

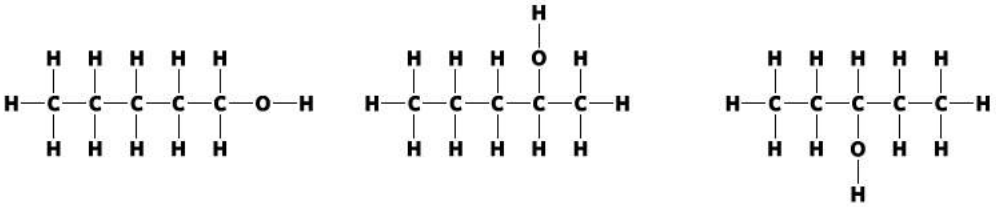


Figure 6 Pentan-1-ol, pentan-2-ol, pentan-3-ol.

The molecules shown in figure 6 all have the same number and types of atoms, and therefore the same molecular formula: $C_5H_{12}O$. However for each compound, the hydroxyl group is bonded to the first, second and third carbon atoms respectively. This leads to molecules which are isomers of each other. More specifically, pentan-1-ol, pentan-2-ol and pentan-3-ol, as shown in figure 6, are **structural isomers** of each other because they differ with respect to the position of the hydroxyl group. This isomerism of alcohols gives rise to diversity amongst the molecules, and different isomers will interact with their surroundings in different ways.

Naming alcohols 1.2.8.9

OVERVIEW

As alcohols are structurally unique from other hydrocarbon compounds, they have their own naming conventions.



Alcohols are named by removing the 'e' from the parent alkane and replacing it with 'ol'.

THEORY DETAILS

When naming an alcohol, the 'e' at the end of the alkane's name is removed and is replaced by 'ol'.

For example, a saturated hydrocarbon with three carbon atoms has the parent alkane name of 'propane'. If there was a hydroxyl group attached to this hydrocarbon, making it an alcohol, the 'e' at the end of 'propane' will be replaced with 'ol', leading to the name 'propanol'. However, for alcohols with more than two carbon atoms it is also important to determine which carbon atom the alcohol functional group is bonded to. As we have done previously with other functional groups, the position of the functional group, which in this case is the hydroxyl group, is signified by placing the carbon number it is attached to, in hyphens, between the parent alkane's name and the 'ol'. Using the example mentioned, if the hydroxyl group was bonded to the first carbon atom, this molecule would be 'propan-1-ol', whereas if it were bonded to the second carbon atom it would be 'propan-2-ol'. These molecules are shown in figure 7.

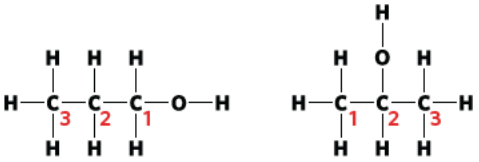
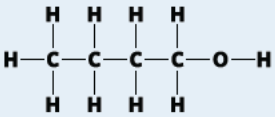


Figure 7 Structural diagram of propan-1-ol (left) and propan-2-ol (right).

As always, we want to number the carbon chain such that the functional group occurs at the lowest numbered carbon. In figure 7, the hydroxyl group could be attached to the 1st carbon atom (if counting from right to left) or the 3rd carbon atom (if counting from left to right). Since only the hydroxyl group is present, we want it to have the lowest numbered position possible and therefore, the molecule is called propan-1-ol and not propan-3-ol.

Worked example 1

What is the systematic name of the molecule with the following structural formula?



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

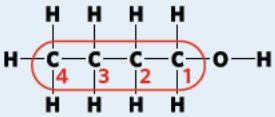
Give the systematic name of the molecule provided.

What strategy(ies) do we need in order to answer the question?

1. Determine the longest carbon chain length.
2. Identify any functional groups.
3. Apply this information using correct IUPAC naming conventions.

Answer

This parent chain contains four carbon atoms, all of which are bonded by single bonds.



Numbering the carbons from right to left allows the hydroxyl group to occur at the lowest number carbon.

The parent alkane name butan- is used for a 4 carbon chain. Additionally, the suffix -1-ol is used to indicate a hydroxyl group on the 1st carbon atom.

The name of this molecule is butan-1-ol.

Carboxylic acids 1.2.8.10

OVERVIEW

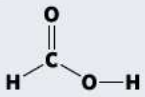
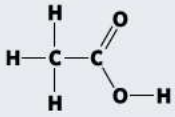
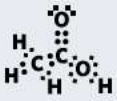
Carboxylic acids are organic compounds that possess a carboxyl functional group.

THEORY DETAILS

A **carboxylic acid** is an organic compound that contains a **carboxyl functional group**, -COOH. This functional group consists of a hydroxyl group, -OH, as well as a carbonyl group, which is a carbon atom double bonded to an oxygen atom, depicted as -C=O. It is important to understand that in carboxyl groups, the carbonyl and hydroxyl groups are considered to be a single group. We do not refer to them separately.

Figure 8 and table 2 show the structural formulas of carboxylic acids.

Table 2 Representations of carboxylic acids.

Name	Molecular formula	Structural formula	Semi-structural formula	Lewis structure
Methanoic acid	CH_2O_2		HCOOH	
Ethanoic acid	$C_2H_4O_2$		CH_3COOH	

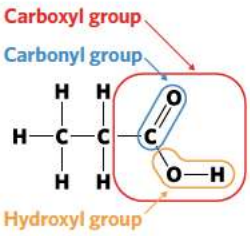


Figure 8 Propanoic acid.

Properties and uses of carboxylic acids

When comparing the electronegativities of the elements found in both the C=O bond and O-H bond, we can identify that these bonds are both polar. As a consequence, carboxylic acids are able to form strong dipole-dipole interactions and hydrogen bonds with other carboxylic acids and other polar molecules such as water. As seen in figure 9, the oxygen atom in the C=O bond of one carboxylic acid is able to form a hydrogen bond with the hydrogen atom from the O-H bond of another carboxylic acid to form a complex known as a **dimer**. A dimer is a molecular complex consisting of two molecules linked together. The formation of dimers between carboxylic acids is one of the reasons why the boiling point of these molecules is higher than alcohols, as the strength of the intermolecular bonding takes a substantial amount of energy to overcome.

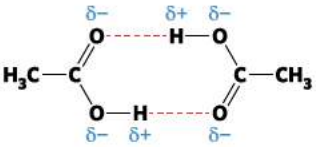


Figure 9 Hydrogen bonding between two carboxylic acids, forming a dimer.

Due to the presence of an additional carbonyl group in carboxylic acids and the ability to form dimers, the boiling points of these molecules are even higher than those of alcohols and alkanes with the same carbon-chain length. Additionally, since carboxylic acids are polar molecules, they can form strong dipole-dipole interactions and hydrogen bonds with water as shown in figure 10. Therefore, these molecules are able to dissolve in water.

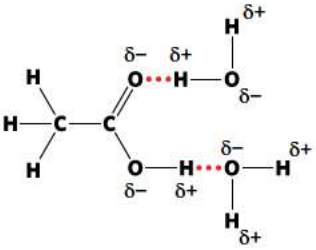


Figure 10 Ethanoic acid forming hydrogen bonds with two water molecules.

However, much like with alcohols, as the carbon-chain length of the carboxylic acid increases, the effect that the functional group has on the properties of the molecule overall is reduced, since the molecule becomes less polar.

It should also be noted that carboxylic acids are weak acids. The concept of acids will be discussed in a later chapter.

Naming carboxylic acids 1.2.8.11

OVERVIEW

Carboxylic acids possess unique naming conventions due to their unique chemical structure.

THEORY DETAILS

When naming carboxylic acids, the parent name of the molecule remains the same as its corresponding alkane. The 'e' at the end of the name of the hydrocarbon is removed and replaced with the suffix '-oic acid'. For example, figure 11 shows a carboxylic acid.

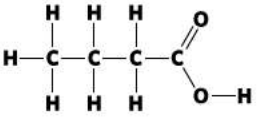


Figure 11 Structural formula of butanoic acid.

As we can see, the parent chain contains four carbons and consists of only carbon to carbon single bonds. The parent name of this molecule is therefore butane. Due to the presence of the carboxyl group, the 'e' is removed, and replaced with 'oic acid' so that the name of the compound is butanoic acid. It's important to note that when determining the parent chain, the carbon involved in the carboxyl group is always included.

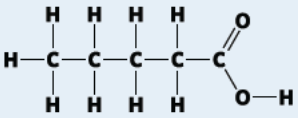
It should also be noted that the carbon atom with carboxylic acid bonded to it is always considered the first carbon atom on the chain and so it is not required to specify the position of the carboxyl group.



Don't forget to include the carbon in the -COOH group when identifying the parent chain of carboxylic acids!

Worked example 2

What is the systematic name of the molecule with the following structural formula?



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

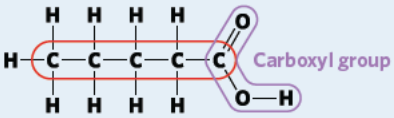
Give the systematic name of the molecule provided.

What strategy(ies) do we need in order to answer the question?

1. Determine the longest carbon chain length.
2. Identify any functional groups.
3. Apply this information using correct IUPAC naming conventions.

Answer

This molecule's parent chain contains five carbon atoms and it is a saturated hydrocarbon.



It has a carboxyl group.

The parent name pentan- is used for a 5 carbon chain. The suffix -oic acid is used to indicate a carboxyl group on the 1st carbon atom.

Final answer: pentanoic acid

Theory summary

- Functional groups are a set of atoms that provide a molecule with a characteristic set of properties.
- Alcohols are organic compounds with a hydroxyl functional group (-OH).
  - Alcohols form strong intermolecular bonds due to the polar hydroxyl group, leading to the properties of a high boiling point and solubility in water.
  - The influence of the hydroxyl group diminishes as the carbon chain length increases.
  - Naming involves identifying the parent chain and adding the suffix '-ol'.
  - The position of hydroxyl group is identified in the name of the compound.
- Carboxylic acids are organic compounds with a carboxyl functional group (-COOH).
  - Carboxylic acids are named through the use of the suffix '-oic acid'.
  - Carboxylic acids are able to form very strong intermolecular bonds due to the polar carboxyl group.
  - The influence of the carboxyl group diminishes as the carbon chain length increases.
  - Carboxylic acids are able to form dimers and very strong hydrogen bonds with each other.

6E QUESTIONS

Theory review questions

Question 1

The functional group of an alcohol consists of oxygen and hydrogen.

A True

B False

Question 2

The name of the alcohol with two carbon atoms is

A ethanol.

B ethanoic acid.

Question 3

The name of a carboxylic acid with three carbon atoms is

A propanoic acid.

B propoxylic acid.

Question 4

A dimer is formed between two carboxylic acids through

A covalent bonding.

B hydrogen bonding.

Question 5

Which of the following is a pair of structural isomers?

A Butan-2-ol, pentan-2-ol

B Pentan-1-ol, pentan-2-ol

Question 6

As the length of a carbon chain increases, the influence of a functional group on the boiling point

A increases.

B decreases.

Question 7

The name of an alcohol with six carbon atoms may be

A hexan-2-ol.

B hexan-5-ol.

Question 8

The molecules commonly used as fuels are

A carboxylic acids.

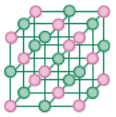
B alcohols.

Question 9

Alcohols have higher boiling points than carboxylic acids with the same number of carbon atoms.

A True

B False



Question 10

A dimer is

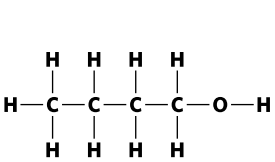
A a complex formed through the intermolecular forces between two carboxylic acid molecules.

B the product of carboxylic acid hydrogen bonding with water.

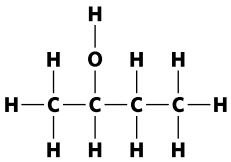
Deconstructed questions

Use the following information to answer questions 11–13.

Timothée has recently learned about alcohols in chemistry class, and in his excitement has drawn some structural formulas and names them to test his understanding.



Butan-4-ol



Butan-2-ol

Question 11 (1 MARK)

The semi-structural formulas of Timothée's molecules, respectively, are

A $CH_3CH_2CH_2CH_2OH$ and $CH_3CH(OH)CH_2CH_3$.

B $CH_3CH_2CH_2OH$ and $CH_3CH(OH)CH_2CH_3$.

C C_4H_9OH for both.

D $CH_3CH_2CH(OH)CH_3$ and $CH_3CH(OH)CH_2CH_3$.

Question 12 (1 MARK)

The molecules proposed by Timothée are

A atomic isomers.

B spatial isotopes.

C molecular isotopes.

D structural isomers.

Question 13 (3 MARKS)

Are the names of the molecules proposed by Timothée correct? Explain your answer, and if either of the proposed names are incorrect, provide the correct names.

Exam-style questions

Within lesson

Question 14 (8 MARKS)

Jeremy is a chemistry student with a passion for anything related to ants. Driven by curiosity, Jeremy decides to research the composition of ant venom, and finds the molecule to be methanoic acid. This molecule, which is injected at high concentration into the body through an ant stinger, creates a pain response.

a What type of molecule is methanoic acid, and what functional group does it contain? (2 MARKS)

b Draw the structural formula of methanoic acid. (2 MARKS)

c Jeremy was surprised to discover that methanoic acid is found naturally in small quantities in the human body, without causing pain. What implication can be derived from this? (1 MARK)

d Having understood methanoic acid, Jeremy wishes to compare this molecule to another type of molecule he's familiar with.

i If the functional group of methanoic acid were replaced with a hydroxyl group, what type of molecule would it be? (1 MARK)

ii What is the IUPAC name of the molecule identified in part 14di? (1 MARK)

iii Draw the structural formula for the molecule identified in part 14di. (1 MARK)

Question 15 (9 MARKS)

Hydrocarbons play a fundamental role in many aspects of modern life, making up the world around us. The addition of functional groups to these molecules can change the way that these molecules interact with each other, and with the world at large.

a Explain how the addition of a hydroxyl group changes the intermolecular forces between hydrocarbons. (3 MARKS)

b Explain how the intermolecular forces between alcohols change the boiling point, compared with hydrocarbons not possessing functional groups. (2 MARKS)

c Carboxylic acids, like alcohols, possess functional groups which enable them to form stronger intermolecular forces.

i What is the name of the functional group found in carboxylic acids? (1 MARK)

ii Describe the structure of a carboxylic acid dimer, including the bonds involved. (3 MARKS)

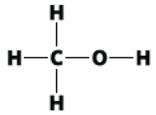
Multiple lessons

Question 16 (11 MARKS)

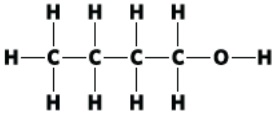
Eileen is a Year 11 chemistry student practicing interpretation of IUPAC nomenclature of alcohols and carboxylic acids in preparation for an upcoming assessment. She comes across a series of structural formulas, and requests some assistance with naming the molecules.

a Name each of the following molecules using IUPAC nomenclature.

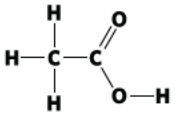
i (1 MARK)



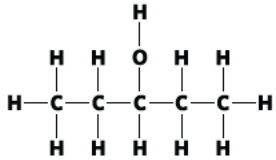
ii (1 MARK)



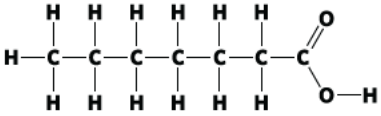
iii (1 MARK)



iv (1 MARK)

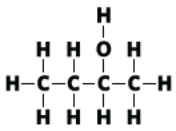
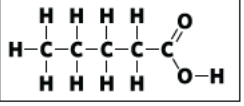


v (1 MARK)



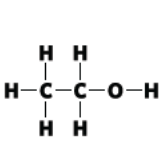
In the course of her studies, she has come across a number of names which she has been unfamiliar with, and filled in her knowledge in the form of a table.

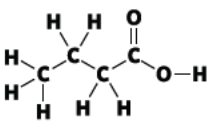
b Fill in the partially completed table below. (6 MARKS)

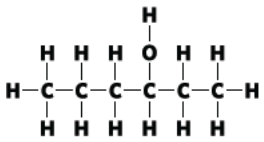
IUPAC Name	Semi-structural formula	Structural formula
Ethanol		
	$CH_3CH_2CH(OH)CH_3$	
	CH_3CH_2COOH	
Pentanoic acid		

Question 17 (8 MARKS)

A university professor is assessing students' understanding of the IUPAC systematic naming conventions, and presents them with the following molecules:







a Provide the IUPAC systematic names of each of these molecules. (3 MARKS)

b Provide the molecular formulas and semi-structural formulas of each of these molecules. (3 MARKS)

c The professor indicates that the carboxylic acid is able to mix in water, which requires overcoming the intermolecular forces of water. Explain why this is the case, and what feature of the carboxylic acid enables this. (2 MARKS)

Question 18 (13 MARKS)

A scientist is investigating the structures and properties of a number of hydrocarbons containing functional groups.

a Draw the structural formulas of the following molecules based on their IUPAC systematic names:

i Pentan-1-ol (1 MARK) **ii** Pentanoic acid (1 MARK) **iii** Butanoic acid (1 MARK)

b Which of these molecules would have the highest boiling point? Explain your reasoning. (4 MARKS)

c Which of these molecules has the greatest molar mass? (1 MARK)

d When comparing each of these molecules to their alkane counterparts, the professor noticed that the molecules possessing functional groups each had a higher boiling point than alkanes with the same parent chain but without functional groups.

i Why do each of the molecules in part a have a higher boiling point than their alkane counterparts? (2 MARKS)

ii Explain why the difference in boiling points between an alcohol and its alkane counterpart decreases as the carbon chain length of the molecules increase. (3 MARKS)

Question 19 (6 MARKS)

Hydrocarbons are a common fuel source used to supply many aspects of modern society. Processes such as fractional distillation are used to separate hydrocarbons based on their boiling point. This same principle applies to the presence of functional groups.

a Compare the intermolecular forces of octane with octan-1-ol. Using this information, determine which molecule has a higher boiling point. (3 MARKS)

b E10 fuel is a commonly used type of fuel, consisting 10% of ethanol and 90% of petrol, which is a mixture of hydrocarbons. This fuel is commonly used due to the fact that ethanol can be produced as a renewable fuel, bioethanol, thus reducing its carbon footprint. However, the percentage of ethanol in the fuel is not increased above 10% due to the engine's inability to function with excess water present.

i A renewable fuel is defined as 'a fuel that can be produced at a rate equal to or greater than its rate of consumption'. Provide one example of why renewable fuels may be beneficial. (1 MARK)

ii Engines are sometimes found to have moisture in them, especially when ethanol is used as a fuel. Using your knowledge about the intermolecular forces formed between ethanol and water, explain why increased concentration of ethanol may increase the amount of water in the engine. (2 MARKS)

Key science skills

Question 20 (8 MARKS)

Eleanor is performing an experiment in which she compares the properties of alcohols and carboxylic acids. In order to do this, Eleanor compared the boiling points of each of these substances by placing them in a test tube and exposing them to a direct flame from a bunsen burner.

a Suggest at least two factors that must be taken into account for this experiment to be carried out safely. (2 MARKS)

b What is the independent variable in this experiment? Explain your answer. (2 MARKS)

c What is the dependent variable in this experiment? (1 MARK)

d What are some variables that must be controlled for the result of this experiment to be valid? Suggest at least three variables. (3 MARKS)

Exam-style question hints

14a: Use numbers for functional groups where appropriate. **14b:** Use the depictions already found in the table as a guide. **14c:** Acids can cause damage to the body. **14di:** Different functional groups give rise to different names. **14dii:** Alcohols contain a -OH functional group. **14diii:** Structural formulas show all covalent bonds. **15a:** Functional groups affect the intermolecular forces between molecules. **15b:** Functional groups can enable different types of bonding. **15c:** Carboxylic acid dimers result from intermolecular interactions. **16a:** Different functional groups give rise to different names. **16b:** The name of organic compounds is dependent upon their structure. **17a:** IUPAC nomenclature has specific suffixes used to indicate particular functional groups. **17b:** Molecular formulas indicate atomic makeup, while semi-structural formulas indicate structure. **17c:** Mixing a substance in water requires overcoming the intermolecular forces in water. **18a:** The names of organic compounds indicate their overall structure. **18b:** Different intermolecular bonds require different amounts of energy in order to break them. **18c:** Molar mass is equal to the molecular weight of a molecule. **18di:** Functional groups affect the boiling point of molecules. **18dii:** The type and magnitude of intermolecular bonds present can affect the boiling point. **19a:** Different intermolecular bonds require different amounts of energy in order to break them. **19bi:** The Earth will run out of non-renewable energy sources at some point in the future. **19bii:** Alcohols possess a hydroxyl functional group. **19biii:** The hydroxyl group has a high affinity for water molecules. **20a:** Safety protocol minimises exposure to hazardous chemicals or flames. **20b:** The independent variable is changed by the experimenter. **20c:** Changes to the dependent variable are observed to determine the effect of changing the independent variable. **20d:** Controlled substances should remain constant and not change experimental conditions between trials.

6F FUNCTIONAL GROUPS - ESTERS

In this lesson, we will be looking at the ester functional group, formed from the reaction of an alcohol and a carboxylic acid.

6A Carbon materials	6B Introduction to hydrocarbons	6C Naming simple hydrocarbons	6D Naming complex hydrocarbons	6E Functional groups - alcohols and carboxylic acids	6F Functional groups - esters	6G Polymers	6H Modifying polymers

Study design dot point

- the grouping of hydrocarbon compounds into families (alkanes, alkenes, alkynes, alcohols, carboxylic acids and non-branched esters) based upon similarities in their physical and chemical properties including general formulas, their representations (structural formulas, condensed formulas, Lewis structures), naming according to IUPAC systematic nomenclature (limited to non-cyclic compounds up to C10, and structural isomers up to C7) and uses based upon properties

Key knowledge units

Esters	1.2.8.12
Naming esters	1.2.8.13

Key terms and definitions

Ester organic compound that contains an ester functional group (-COO-)

Catalyst substance that increases the reaction rate without itself being consumed or permanently changed

Esterification reaction chemical reaction between organic compounds that forms at least one ester as a product

Condensation reaction chemical reaction where two molecules combine with the release of a water molecule

Hydrolysis reaction chemical reaction where water is used to break the bonds of a substance

Lesson links

This lesson builds on:

▶ 5C Intermolecular bonding

Functional groups alter the intermolecular interactions of molecules.

▶ 6C Naming simple hydrocarbons

Hydrocarbons can contain different types of functional groups.

Esters 1.2.8.12

OVERVIEW

Esters are organic compounds consisting of two carbon chains separated by an ester group.

THEORY DETAILS

Esters are another family of chemical compounds derived from hydrocarbons. They are defined by the -COO- functional group which is referred to as an ester group.

An ester functional group is similar to a carboxyl functional group, consisting of a carbonyl group (C=O) with a second oxygen atom covalently bonded to the carbonyl carbon atom. This can be represented as -COO-, and an example of this is shown in figure 1. The semi-structural formula of this ester is $CH_3CH_2COOCH_3$, with the presence of the ester functional group shown by the -COO- in the middle of the semi-structural formula.

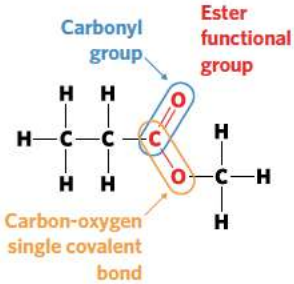


Figure 1 Structural formula of an ester.

At first glance, the ester functional group and the carboxyl functional group seem quite similar. In fact, esters are produced through a reaction between a carboxylic acid and an alcohol in the presence of concentrated sulfuric acid. The sulfuric acid acts as a **catalyst** which is a substance that is not consumed in a chemical reaction but increases the rate of reaction. The catalyst is written above the reaction arrow. This process is called an **esterification reaction**, and occurs between the carboxyl functional group and the hydroxyl functional group of the carboxylic acid and alcohol respectively. For VCE Chemistry, the alcohol will have the hydroxyl group (-OH) at the end of the molecule. As shown in figure 2, water is also produced during this process and therefore this reaction can also be referred to as a **condensation reaction**.

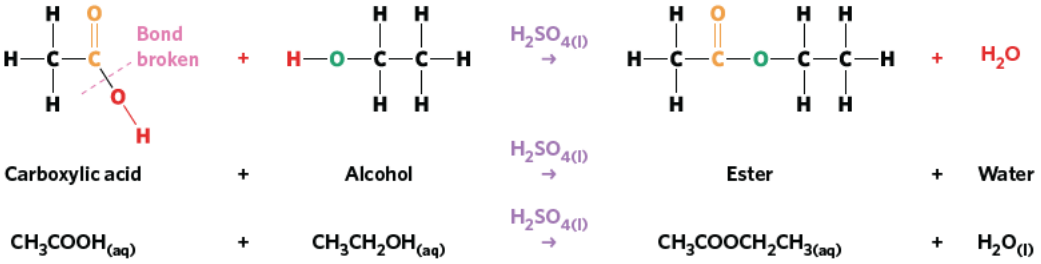


Figure 2 Esterification reaction between ethanoic acid and ethanol to produce ethyl ethanoate and water.

The hydrogen atom from the hydroxyl group of the alcohol combines with the hydroxyl group from the carboxylic acid to form water, which is eliminated from the reaction. The remaining part of the alcohol and the carboxylic acid are bonded to each other to form the ester.

Considering water is released during esterification, water is therefore required as part of the reaction when esters are broken down into their original alcohol and carboxylic acid. This process is called a **hydrolysis reaction**, as shown in figure 3. Through the input of water, the covalent bond between the carboxylic acid residue and the alcohol residue is broken, and the atoms in the water molecule are used to reform the hydroxyl and carboxyl functional groups into their original forms, as highlighted in red and blue in figure 3.

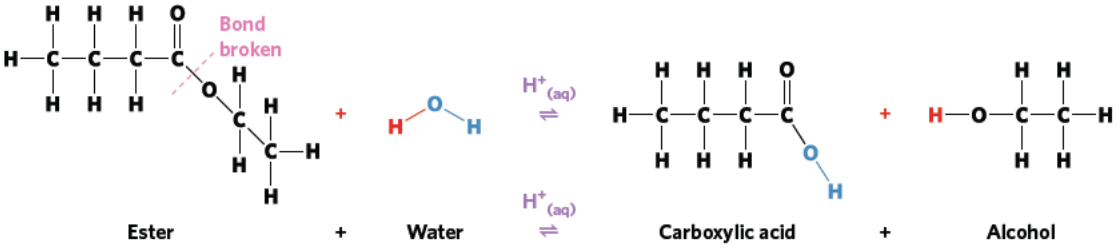


Figure 3 Hydrolysis reaction of an ester.

Worked example 1

Write the equation using structural formulas for the chemical reaction between propanoic acid and methanol. Include, if present, any catalysts.

What information is presented in the question?

The reactants are propanoic acid (a carboxylic acid) and methanol (an alcohol).

What is the question asking us to do?

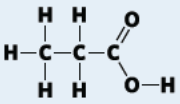
Represent the chemical reaction that occurs (an esterification reaction) by using the structural formulas of the reactants and products.

What strategy(ies) do we need in order to answer the question?

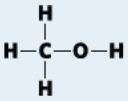
1. Determine the structural formulas for propanoic acid and methanol.
2. Determine the structural formulas for the products, an ester and a water molecule.
3. Write out the chemical equation, including the concentrated sulfuric acid.

Answer

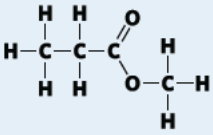
Propanoic acid is a carboxylic acid with three carbon atoms and so has the following structural formula:



Methanol is an alcohol with one carbon atom and so has the following structural formula:

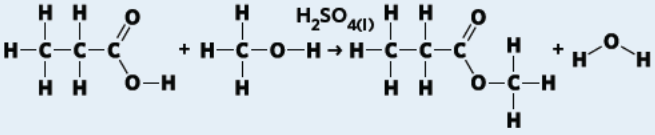


An esterification reaction is between the carboxyl (-COOH) and hydroxyl group (-OH) of the carboxylic acid and alcohol respectively where the carboxyl group loses an OH group and the alcohol loses an H atom to form the ester group -COO-.



A water molecule is produced as part of the reaction. The reaction also requires a concentrated sulfuric acid catalyst.

With all this information, the equation for this reaction is:



Properties and uses of esters

Similar to a carboxylic acid, esters contain a polar C=O bond. As a result, esters are able to take part in permanent dipole-dipole interactions with other polar compounds. However, compared to carboxylic acids, the ester functional group is unable to form hydrogen bonds with itself and therefore has a lower boiling point. This is summarised in table 1 which shows the boiling points of four organic compounds, each with four carbon atoms, with different functional groups.

Table 1 Boiling point comparison.

	Butane (alkane)	Butan-1-ol (alcohol)	Butanoic acid (carboxylic acid)	Ethyl ethanoate (ester)
Boiling point at atmospheric pressure	−0.5°C	117.7°C	163.7°C	77.1°C

However, although esters are unable to form hydrogen bonds with other esters, they can form hydrogen bonds with water as shown in figure 4.

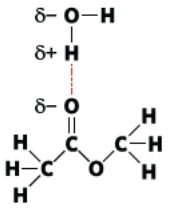


Figure 4 Hydrogen bonding between water and an ester.

As a result of this ability to form hydrogen bonds, small esters are able to dissolve in water but its ability to dissolve in water decreases as the chain length of the ester increases. This is because of the decreasing importance of the functional group compared to the larger molecule, as explained in lesson 6E.

Esters are characterised by their sweet and fruity odours, and are the main chemicals that produce the distinctive smells of fruits. Consequently, many products such as perfumes, food flavourings and essential oils utilise esters to produce their scent. In addition to this, some particularly volatile esters are used as solvents for lacquers, paints and varnishes.

Naming esters 1.2.8.13

OVERVIEW

The unique structure of esters means that there are unique naming conventions.

THEORY DETAILS

Naming esters is a bit more complicated than other hydrocarbons, primarily due to the fact that esters originate from two different types of organic compounds. However, the same principles apply as before, except now the carbon chain length of both carbon chains must be accounted for.

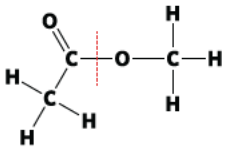


Figure 5 Methyl ethanoate.

When naming an ester, it is important to distinguish where each of the carbon chains is bonded, as one of the chains will be associated with the carbonyl group and the other will be bonded to the single oxygen atom.

Looking at figure 5, we can break down the ester into the two constituent molecules that it is derived from, a carboxylic acid and an alcohol, which come together to form an ester. The component of the ester derived from a carboxylic acid is the carbon chain where the carbonyl group (C=O) can be found, and the component derived from an alcohol is the carbon chain bound to the single oxygen atom as shown in figure 6.

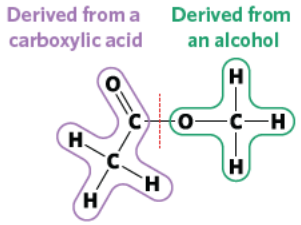


Figure 6 Breakdown of an ester.

The component derived from an alcohol in this case can be seen to have a single carbon atom in its carbon chain. When naming an ester, the component derived from an alcohol is named first, with its parent alkane name followed by the suffix '-yl' after the 'ane' has been removed. Therefore, we can name this component 'methyl'.

When looking at the carbon chain derived from a carboxylic acid (the chain with the carbonyl group), we can see that the structure contains two carbons. We use the parent alkane name after the 'e' has been removed and in order to indicate that this is the part of the ester that is derived from a carboxylic acid, we use the suffix '-oate'. Therefore, we can determine that the second part of the name of this ester is 'ethanoate'.

Putting this all together, the name of this molecule is methyl ethanoate as shown in figure 7.

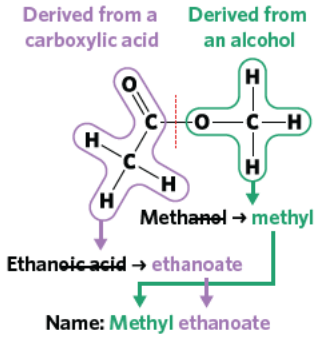
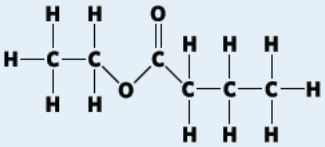


Figure 7 The process of naming an ester.

Worked example 2

What is the systematic name of the molecule with the following structural formula?



What information is presented in the question?

The structural formula for a molecule.

What is the question asking us to do?

Give the systematic name of the molecule provided.

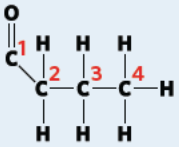
What strategy(ies) do we need in order to answer the question?

1. Determine the carbon chain lengths on either side of the ester group.
2. Identify the parent names of the carbon chains present.
3. Apply identified parent names of carbon chains in the molecule to IUPAC nomenclature.

Answer

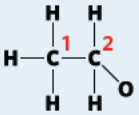
This molecule is an ester due to the presence of the -COO- functional group.

The 4 carbon chain is part of the carbonyl group side.



The parent name butan- is used for the 4 carbon chain, and the suffix -oate is used to indicate it is bound to the carbonyl group.

The 2 carbon chain is bonded to the single oxygen atom in the ester group.



The parent name eth- is used for the 2 carbon chain, and the suffix -yl is used to indicate it is on the side not bound to the carbonyl group.

Final answer: ethyl butanoate

Theory summary

- Esters are organic compounds with an ester functional group (-COO-).
- Esters consist of an alcohol and a carboxylic acid component.
- Esters are produced through an esterification/condensation reaction between a carboxylic acid and an alcohol.
- Esters have a lower boiling point than alcohols and carboxylic acids.
- Esters can be broken down into constituent components through a hydrolysis reaction.
- Esters are named in two parts, the alcohol component is named first with the suffix '-yl', followed by the carboxylic acid component with the suffix '-oate'.
- Esters have a characteristic sweet, fruity odour.

6F QUESTIONS

Theory review questions

Question 1

The ester functional group contains

A hydrogen.

B oxygen.

Question 2

There are only single carbon-oxygen covalent bonds in the ester functional group.

A True

B False

Question 3

Esters are formed in

A condensation reactions.

B hydrolysis reactions.

Question 4

Which of the following options correctly identifies the reactants and products of an esterification reaction?

	Products	Reactants
A	Alcohol, carboxylic acid	Water, ester
B	Water, ester	Alcohol, carboxylic acid

Question 5

A student mixed ethanol and ethanoic acid together without any other chemicals but no chemical reaction occurred. A possible explanation is that

A the alcohol and carboxylic acid are required to have a different number of carbon atoms for a reaction to occur.

B no concentrated sulfuric acid catalyst was present.

Question 6

Esters are characterised by their

A sour, bitter smell.

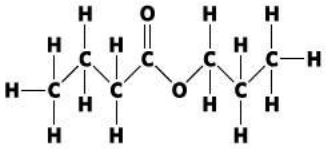
B fruity, sweet smell.

Question 7

The alcohol used to produce this ester

A contains four carbons in the parent chain.

B contains three carbons in the parent chain.

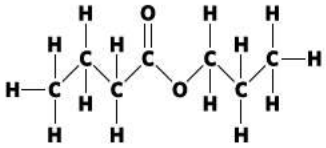


Question 8

The carboxylic acid used to produce this ester

A contains four carbons in the parent chain.

B contains three carbons in the parent chain.



Question 9

Esters are unable to form hydrogen bonds in any circumstances.

A True

B False

Question 10

The boiling point of esters is lower than carboxylic acids with a comparable number of carbon atoms as there are no free hydroxyl groups for hydrogen bonding.

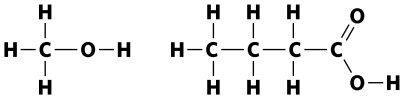
A True

B False

Deconstructed questions

Use the following information to answer questions 11–13.

Jeremy was asked to conduct an experiment using the molecules shown:



As the experiment proceeded, he noticed that there was an increase in the volume of water, and also that there was a fruity smell coming from the test tube.

Question 11 (1 MARK)

What type of reaction would take place between these two molecules?

A Hydrolysis
B Substitution
C Condensation
D Hydrogenation

Question 12 (1 MARK)

What is the functional group that would be present in the product?

A Carboxyl group
B Hydroxyl group
C Ester group
D Ethyl group

Question 13 (2 MARKS)

Explain the phenomena of the increasing volume of water and the fruity smell coming from the test tube.

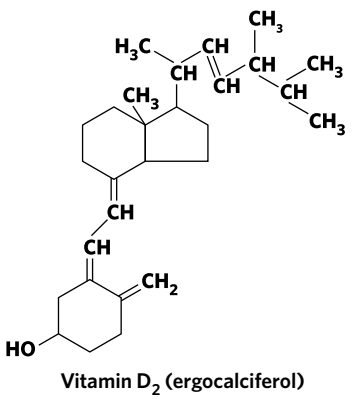
Exam-style questions

Within lesson

Question 14 (1 MARK)

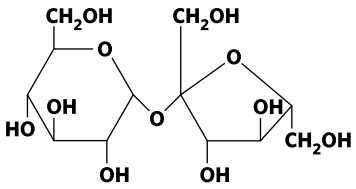
Which of the following biomolecules contains an ester linkage?

A



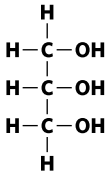
Vitamin D_2 (ergocalciferol)

B



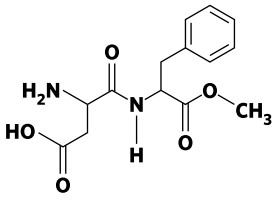
Sucrose

C



Glycerol

D



Aspartame

Question 15 (10 MARKS)

There are many different types of esters that exist that are responsible for a multitude of different smells, ranging from banana to pineapple.

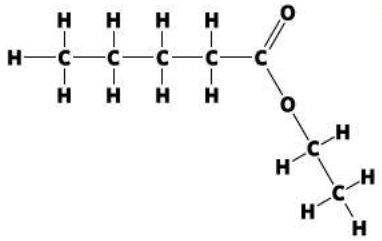
a Name the ester produced from the reactions of the following carboxylic acid and alcohol.

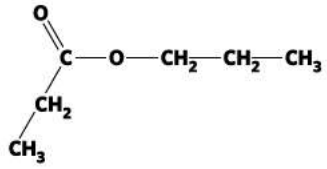
i Butan-1-ol and methanoic acid (1 MARK)

ii Propan-1-ol and pentanoic acid (1 MARK)

iii Hexan-1-ol and butanoic acid (1 MARK)

b Give the systematic name for each ester presented.

i  (1 MARK)

ii  (1 MARK)

iii $CH_3OOCCH_2CH_2CH_2CH_2CH_3$ (1 MARK)

iv $CH_3CH_2COOCH_3$ (1 MARK)

c Draw the structural formula for each of the following molecules.

i Methyl ethanoate (1 MARK)

ii Propyl butanoate (1 MARK)

iii Butyl ethanoate (1 MARK)

Multiple lessons

Question 16 (8 MARKS)

Biodiesel is a renewable fuel which consists of esters (specifically, fatty acid methyl esters). Diesel is a non-renewable fuel which consists of straight-chain alkanes. Hygroscopicity is the property of fuels to absorb water. The most desirable fuel is the one that is able to flow along fuel lines easily and not absorb water.

a What is the difference between diesel and biodiesel with respect to the functional groups present? (2 MARKS)

b As a result of this difference in functional group(s), compare the melting points of diesel and biodiesel. Which fuel would be more suitable in colder climates? (3 MARKS)

c As a result of this difference in functional group(s), compare the hygroscopicity of diesel and biodiesel. Which fuel would be more suitable in wet, humid conditions? (3 MARKS)

Question 17 (8 MARKS)

Esterification reactions produce ester molecules, which possess a characteristically sweet, fruity smell.

a Write a worded reaction showing the reactants, products and necessary chemicals in an esterification reaction. (2 MARKS)

b What type of reaction is an esterification reaction? (1 MARK)

c A scientist chooses to use ethanol as the alcohol in an esterification reaction. Draw the structural formula of ethanol. (1 MARK)

d The ethanol reacts with propanoic acid. Draw the structural formula of propanoic acid. (1 MARK)

e Draw the structural formula of the largest product of the esterification reaction. Circle and name the functional group present. (2 MARKS)

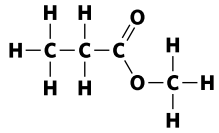
f What is the name of the ester produced? (1 MARK)

Question 18 (13 MARKS)

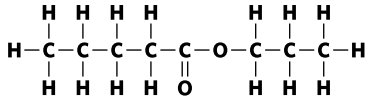
The science department at a university possesses a wide array of esters for students to practice identification and to undertake experiments with them.

a Name each ester shown and identify the carboxylic acid and alcohol used.

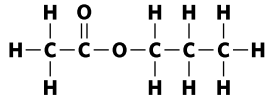
i (3 MARKS)



ii (3 MARKS)



iii (3 MARKS)



b Why is water required as an input in the hydrolysis of esters? (2 MARKS)

c Why is concentrated sulfuric acid required in esterification reactions? (2 MARKS)

Question 19 (13 MARKS)

Jeremy decides to conduct his own set of experiments to determine what substances can be produced using methanoic acid. He possesses a bottle of concentrated methanoic acid, as well as some other common substances, such as ethanol. As part of his testing, Jeremy mixes methanoic acid and ethanol together in the presence of heat and acid.

a What type of reaction takes place between methanoic acid and ethanol? (1 MARK)

b Draw the reaction including the structural formulas of all reactants and products. (3 MARKS)

c What are the names of the products produced in this reaction? (2 MARKS)

d Explain the presence of the byproduct. (2 MARKS)

e Suggest an observation that will signal the formation of an ester. (1 MARK)

f Compare the boiling points of ethanol, ethanoic acid and methyl methanoate. Suggest how this property could be used to extract the ester from a mixture containing an ester, alcohol and carboxylic acid. (4 MARKS)

Key science skills

Question 20 (5 MARKS)

A student is experimenting with esterification reactions in their school laboratory. She recalls from her studies of esters that these molecules have a sweet, fruity smell, and decides to directly smell the tube to assess whether the reaction had taken place.

a Why is this practice dangerous? How could the student improve the safety of her assessment methods? (2 MARKS)

b When measuring the amount of ester formed, the student accidentally used a set of scales that was incorrectly calibrated throughout the entire experiment.

i What type of error is this? (1 MARK)

ii What is the definition of this error type? (1 MARK)

iii How can this error be corrected when viewing the results? (1 MARK)

Exam-style question hints

14: Different functional groups contain different types of atoms and different types of bonds. **15a:** The name of an ester depends on the name of the carboxylic acid and alcohol. **15b:** The name of an ester depends on the length of the carbon chains on either side of the ester functional group. **15c:** The carbonyl component of the ester functional group is derived from the carboxylic acid. **16a:** The ester functional group contains a carbon-oxygen double bond and a carbon-oxygen single covalent bond. **16b:** The ester functional group is polar and can form permanent dipole-dipole bonds with itself. **16c:** The ester functional group can form hydrogen bonds with water. **17a:** Esterification reactions require a catalyst. **17b:** Esterification reactions produce a water molecule. **17c:** Alcohols have a hydroxyl group (-OH). **17d:** Carboxylic acids contain a carboxyl group (-COOH). **17e:** The products of an esterification reaction are an ester and a water molecule. **17f:** The carbon chain derived from the alcohol is named before the carbon chain derived from the carboxylic acid. **18a:** The carbon chain derived from the alcohol is named before the carbon chain derived from the carboxylic acid. **18b:** Water is a product in esterification reactions. **18c:** Catalysts help reactions to occur faster. **19a:** The reaction between a carboxylic acid and alcohol results in the production of water. **19b:** Structural formulas show all covalent bonds and elements involved. **19c:** The carbon chain derived from the alcohol is named before the carbon chain derived from the carboxylic acid. **19d:** Byproducts are chemicals that are produced in reaction but are not the desired product. **19e:** Chemical reactions can be observed through the senses like sight and smell. **19f:** Compounds can be separated based on differences in boiling points. **20a:** Volatile organic compounds can be toxic when inhaled. **20b:** Errors affect results in different ways.

6G POLYMERS

In this lesson, we will be discussing the formation and properties of polymers, as well as the roles that these molecules play in contemporary society.

6A Carbon materials | 6B Introduction to hydrocarbons | 6C Naming simple hydrocarbons | 6D Naming complex hydrocarbons | 6E Functional groups - alcohols and carboxylic acids | 6F Functional groups - esters | 6G Polymers | 6H Modifying polymers

Study design dot points

- the formation of polymers from monomers including addition polymerisation of alkenes
- the distinction between linear (thermoplastic) and cross-linked (thermosetting) polymers with reference to structure, bonding and properties including capacity to be recycled

Key knowledge units

Formation of polymers	1.2.10.1
Thermoplastic vs thermosetting polymers	1.2.11.1

Key terms and definitions

Polymer molecule produced from the addition reaction between monomers

Addition polymerisation reaction that joins monomers together to form a polymer

Monomer molecules that can react with other molecules to form larger molecules (polymers)

Low density polyethene branched polymer produced at high pressures

High density polyethene tightly packed polymer produced at low pressures

Thermoplastic polymer polymer with relatively weak intermolecular bonds that can be remoulded at a particular temperature

Thermosetting polymer polymer with relatively strong intermolecular bonds that degrades when heated at a particular temperature and cannot be remoulded

Elastomer polymers that form occasional cross-links and are elastic

Lesson links

This lesson builds on:

- 6A Carbon materials

Polymers are carbon materials.

Formation of polymers 1.2.10.1

OVERVIEW

Polymers are formed through joining together monomers.

THEORY DETAILS

Polymers are carbon materials that are also known as plastics. These materials are found abundantly in society as they are cheap, easy to manufacture, user-friendly and versatile. In this lesson, we will be covering how the formation of polymers takes place, and how this process provides them with these unique properties.

The term **polymer** implies a substance consisting of many smaller molecules, as the prefix 'poly-' means 'many'. A polymer is formed through **addition polymerisation** of smaller molecules, referred to as **monomers** ('mono-' is 'one'), which involves joining many monomers together in a chain. This is demonstrated in figure 1.

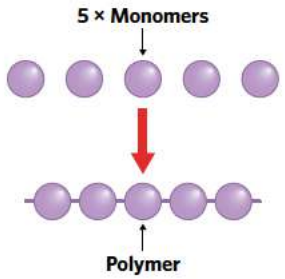


Figure 1 Polymer formation.

Addition polymerisation

As we saw in lesson 6B, hydrocarbons such as alkenes participate in addition reactions, where a bond is broken in a double bond to produce a single bond with another molecule. This occurs because ethene molecules are unsaturated, however the product of the reaction will be a saturated hydrocarbon. This same process takes place in the formation of polymers. One type of polymer is polyethene, produced from the addition polymerisation of ethene molecules, as shown in figure 2. To produce one molecule of polyethene, thousands of ethene molecules must undergo addition reactions. It is also worth noting that any unsaturated molecule can participate in addition polymerisation reactions.

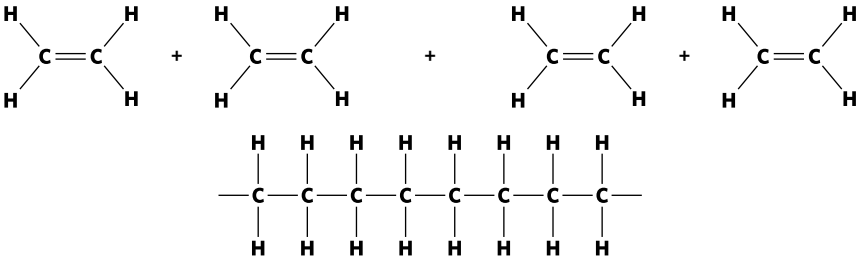


Figure 2 Addition polymerisation of ethene.

The name of a polymer can be determined by adding the prefix 'poly-' before the monomer that it was made from. For example, a polymer produced from propene molecules would be referred to as 'polypropene'.

Figure 2 also demonstrates one possible representation of polymers, with the unconnected ends on either side indicating that the molecule's length continues. However, the most common method of representing polymers is shown in figure 3. Since polymers consisting of only one monomer have a repeating structure, a polymer and its constituent monomer have the same empirical formula.

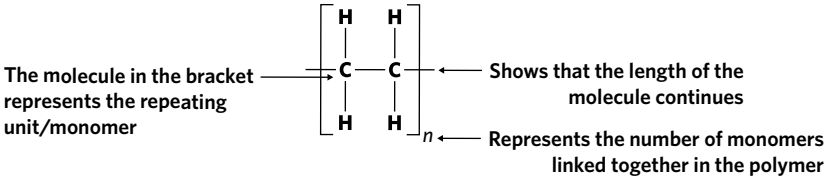


Figure 3 Polymer representation.

While the term plastic is often used to refer to polymers, in chemistry the term is used to describe a property of a material. A plastic material is one that can be moulded into different shapes easily. Objects made out of polymers, such as bottles or washing baskets, are able to hold their unique shape due to the properties of their material, as demonstrated in figure 4.



Image: Ivaschenko Roman, Studio Light and Shade/Shutterstock.com

Figure 4 Plastic bottles and washing basket.

Polymers are hydrocarbons, and therefore their properties will alter in similar ways to hydrocarbons. As we have discussed in previous lessons, as the length of a hydrocarbon increases, the melting point also increases due to increased dispersion forces. Since polymers possess very long chains, these chains have strong intermolecular bonding and, because of this, are found in a solid state at room temperature.

Polymers are highly modifiable and many types of polymers possess their own unique properties. Some general characteristics seen in most polymers include:

- Light and strong
- Durable
- Highly versatile and modifiable – can possess many colours and characteristics
- Flammable
- Good chemical resistance – particularly against acids
- Effective thermal and electrical insulators

Generally, polymers are produced from petroleum, however there has been a recent push to produce polymers from more environmentally-friendly sources due to the impact of petroleum on the environment and of depleting fossil fuel sources. Some companies, such as LEGO®, are now initiating a shift from petroleum-sourced polymers to producing polymers from renewable sources such as sugarcane.

Low density polyethene

When polymerisation takes place at high temperatures and high pressures, these conditions prevent the polymer from forming symmetrical, linear chains. Polyethene chains produced under these conditions are known as **low density polyethene** (LDPE), and possess many branches coming off from the parent chain. Figure 5 demonstrates the atomic and molecular structures of LDPE, as well as an example of LDPE.

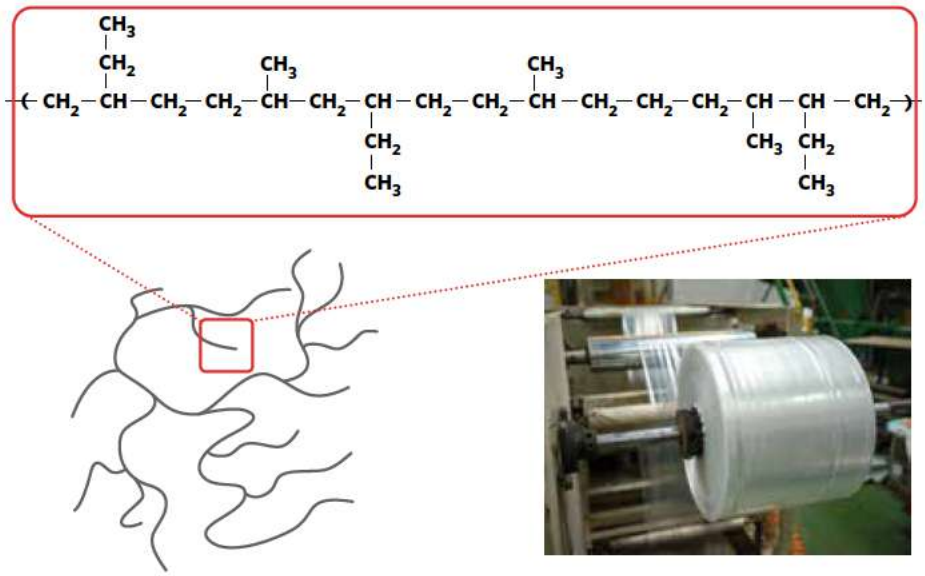


Image: lowpower225/Shutterstock.com

Figure 5 Structures and example of low density polyethene.

The presence of these branches prevents low density polyethene from packing together closely, leading to weakened dispersion forces and a greater amount of separation between molecules. Because of this, low density polyethene possesses the following properties:

- Lower density
- Relatively soft
- Lower melting point (105–115°C)
- Good chemical resistance
- Non-crystalline
- Opaque, but transparent in thin forms
- Insulator of electricity

High density polyethene

In contrast to low density polyethene, **high density polyethene** (HDPE) is produced without the need for high pressures. These conditions lead to the formation of very few branches therefore the molecules are able to pack together tightly. This increases the density and hardness of the product. The properties of high density polyethene include:

- Higher density
- Hard
- Higher melting point (120–180°C)
- Weatherproof and cold-resistant
- Translucent
- Insulator of electricity
- Good chemical resistance
- Crystalline regions

Figure 6 demonstrates the structure and examples of HDPE polymers.

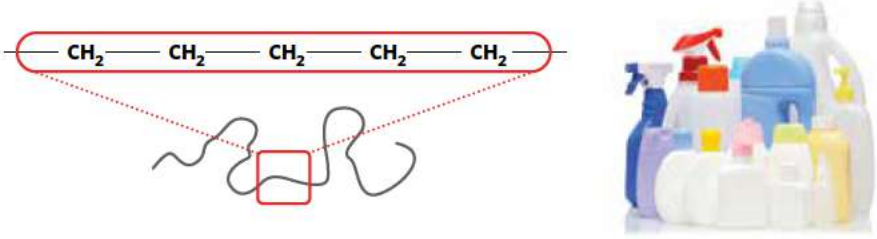


Image: MrJPEG/Shutterstock.com

Figure 6 Structures and example of high density polyethene.

Other addition polymers

Polymers can be used for a wide range of applications, and the structure of these molecules can be tweaked to enhance the molecule for a particular function. For example, the addition of a highly electronegative atom such as chlorine to a monomer will greatly increase the strength of intermolecular forces between polymer chains, thus changing the properties of the substance.

The ability to modify polymers to carry out particular functions will be discussed in greater detail in the next lesson.

Thermoplastic vs thermosetting polymers 1.2.11.1

OVERVIEW

Thermoplastic and thermosetting polymers possess different properties and are processed differently.

THEORY DETAILS

Polymers can be separated into two major categories – **thermoplastic polymers** and **thermosetting polymers**. Polymers are assigned to a category depending on how they interact with heat.

Thermoplastic polymers

When heated, thermoplastic polymers will soften, which enables them to be reshaped or remoulded. The bonds within thermoplastic polymers are stronger than the intermolecular forces holding together the individual molecules (hydrogen bonds, dipole-dipole bonds and dispersion forces), and therefore the intermolecular bonds will break upon heating. This causes molecules to become free and mobile which enables them to be reshaped.

Polymers that are able to be reshaped are more likely to be recyclable. This is an advantageous property of thermoplastic polymers today, where the environmental effects of polymers as well as diminishing fossil fuel supplies are highlighting the importance of recycling. An example of a thermoplastic polymer is Acrylonitrile butadiene styrene (ABS), which is the plastic used in lego. The lego bricks shown in figure 7 can be melted down and remoulded for other purposes.



Image: 3d_kot/Shutterstock.com

Figure 7 Lego bricks.

Thermosetting polymers

When thermosetting polymers are heated, these molecules will burn or decompose. Therefore it is difficult to recycle these polymers.

Thermosetting polymers decompose when exposed to heat because the bonds between polymer chains are stronger than in thermoplastic polymers. Thermosetting polymers often possess covalent bonds between polymer chains and therefore the strength of the intermolecular bonds in these polymers is comparable to the intramolecular bonds. When exposed to heat, the intermolecular bonds are just as likely to break as the intramolecular bonds. Consequently, breaks can take place at any point in the polymer, causing degradation and decomposition of the polymer.

Covalent bonds between polymer chains are referred to as cross-links. The presence of cross-links in a molecule greatly restricts the movement of a molecule, making it more rigid and heat resistant. Due to these properties, cross-setting polymers are used to make tough plastics, such as bowling balls. An example of a thermosetting polymer is silicone, which is used for a number of purposes such as sealants and utensils. The utensils shown in figure 8 therefore cannot be remoulded and recycled.



Image: IMG Stock Studio/Shutterstock.com

Figure 8 Silicone utensils.

Elastomers

The primary structural difference between thermoplastic and thermosetting polymers is the presence of cross-links in thermosetting polymers. Another group of polymers is **elastomers**, which possess cross-links like thermosetting polymers, but only occasionally. Since the intermolecular bonds in elastomers are stronger than those in thermoplastic polymers but weaker than those in thermosetting polymers, elastomers can move past each other when stretched. However, due to the elasticity of the cross-links, elastomers will return back to their original shape, and cannot be remoulded when heated. An example of elastomers is rubber bands, which can be easily stretched but will return back to their original shape as shown in figure 9.



Image: Anton Starikov/Shutterstock.com

Figure 9 An example of elastomers.

Worked example 1

Jeremy is a mechanical engineer, and he suggests that a new washing basket being produced should be made from thermoplastic polymers due to the potential to be recycled.

Explain the difference between thermoplastic and thermosetting polymers, referencing recyclability.

What information is presented in the question?

It has been suggested that thermoplastic polymers should be used to produce a washing basket for recyclability.

What is the question asking us to do?

Explain the difference between thermoplastic and thermosetting polymers.

What strategy(ies) do we need in order to answer the question?

1. Recall the structure of thermoplastic polymers.
2. Identify the properties of thermoplastic polymers.
3. Recall the structure of thermosetting polymers.
4. Identify the properties of thermosetting polymers.

Answer

Thermoplastic polymers possess weaker intermolecular bonds than thermosetting polymers, including hydrogen bonds, dipole-dipole bonds and weak dispersion forces. This means that these bonds are broken when the polymer is heated, enabling remoulding and recycling.

Conversely, thermosetting polymers possess strong covalent intermolecular bonds between polymer chains. This means that when these polymers are heated, the covalent intermolecular bonds have the same chance of being broken as the intramolecular bonds, increasing the likelihood of decomposition and preventing remoulding and recycling.

Theory summary

- Polymers are formed from the addition polymerisation of unsaturated monomers.
- Low density polyethene is produced at higher pressure than high density polyethene and consequently has a less symmetrical structure.
- Thermoplastic polymers have weaker intermolecular bonds and can be remoulded and recycled.
- Thermosetting polymers possess strong cross-links and decompose when exposed to heat.
- Elastomers possess occasional cross links providing them with elasticity.

6G QUESTIONS

Theory review questions

Question 1

Polymers are made up of

A thousands of monomers in a chain.

B a web of monomers.

Question 2

An example of a monomer is

A propane.

B ethene.

Question 3

The process in which polymers are produced is called

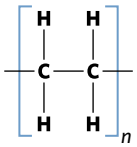
A addition polymerisation.

B reduction polymerisation.

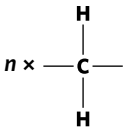
Question 4

Which of the following is an appropriate representation of a polymer?

A



B



Question 5

The main difference in producing low and high density polyethene is that

A low density polyethene is produced at lower pressures.

B low density polyethene is produced at lower temperatures.

Question 6

Thermoplastic polymers

A are able to be remoulded when heated.

B are resistant to remoulding.

Question 7

Thermosetting polymers

A can be recycled.

B are hard and rigid.

Question 8

The type of intermolecular bonding in thermoplastic polymers may be

A covalent bonding.

B dipole-dipole bonding.

Question 9

The name of the bonds formed in thermosetting polymers are

A cross-links.

B joining bonds.

Question 10

Elastomers

A possess occasional covalent intermolecular bonds.

B are able to be permanently remoulded when heated.

Deconstructed questions

Use the following information to answer questions 11–13.

A Year 11 Chemistry student who has recently learned about different types of polymers has set up a series of experiments to determine the difference between thermoplastic and thermosetting polymers.

Question 11 (1 MARK)

Which of the following best describes thermoplastic polymers?

A Polymer with intermolecular forces weaker than intramolecular forces, enabling remoulding upon heating.

B Polymer with strong intermolecular forces, enabling it to not degrade when being remodelled.

C Polymer with covalent intermolecular forces which, when broken, cause the molecule to decompose.

D Polymer with covalent intermolecular forces, meaning that when the molecule is heated either intermolecular bonds or intramolecular bonds may be broken.

Question 12 (1 MARK)

Which of the following best describes thermosetting polymers?

A Polymer with intermolecular forces weaker than intramolecular forces, enabling remoulding upon heating.

B Polymer with strong intermolecular forces, enabling it to not degrade when being remodelled.

C Polymer with covalent intermolecular forces which, when broken, cause the molecule to decompose.

D Polymer with covalent intermolecular forces, meaning that when the molecule is heated either intermolecular bonds or intramolecular bonds may be broken.

Question 13 (2 MARKS)

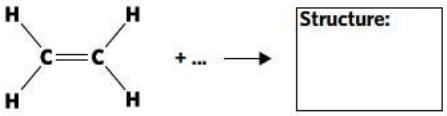
Briefly describe the general structure of polymers, including what they are composed of and how they are formed.

Exam-style questions

Within lesson

Question 14 (3 MARKS)

Please complete the following reaction using structural formulas:



Name: _____ **Polyethene**

Question 15 (1 MARK)

The difference between low density and high density polyethene is:

A HDPE is produced at higher pressures causing compaction.

B LDPE is produced at higher pressures preventing the formation of symmetrical chains.

C HDPE is produced at lower temperatures, allowing more time for symmetrical chains to form.

D LDPE is produced at lower pressures, causing polymer chains to be more dispersed.

Question 16 (1 MARK)

Silicone is a thermosetting polymer. Utensils made out of silicone

A can be heated and remoulded into a new shape.

B cannot be remoulded as the intermolecular forces are too weak to maintain structural integrity.

C can be recycled.

D possess bonds between polymers and within polymers, that are of similar strength.

Question 17 (1 MARK)

Elastomers

A can be stretched but return back to their original shape due to the presence of cross-links.

B can be easily remoulded.

C are rigid structures unable to be stretched.

D have strong ionic bonds between polymer chains.

Question 18 (9 MARKS)

Grant is an engineer who wants to design a new material using polymers. When considering the polymer to use, Grant considers the properties of the polymer, as well as how easily the material can be remoulded and recycled. Having picked a monomer to make the polymer from, Grant has to work out how to carry out the production.

a What is the name of the reaction in which polymers are formed from monomers? (1 MARK)

b When producing the polymer, Grant has the temperature set to a given temperature but is given the choice between a higher and lower pressure setting.

i Identify the name and properties of the polymer that would be produced at a higher pressure setting. (2 MARKS)

ii Identify the name and properties of the polymer that would be produced at a lower pressure setting. (2 MARKS)



c Having formed his polymer, Grant decides to try and heat the polymer up for remoulding. However, upon heating, the polymer decomposes.

i What type of polymer did Grant make, and why did the polymer decompose when heated? (3 MARKS)

ii Identify one property of the polymer produced. (1 MARK)

Multiple lessons

Question 19 (9 MARKS)

Edward spent the last number of weeks studying carbon materials, including polymers. Knowing that polymers are formed from monomers, Edward sets out to determine which monomers can be used to form a polymer.

a Identify which of the following could be used to produce a polymer. Explain your answer for the monomers that cannot be used. (5 MARKS)

I Ethene **II** Methane **III** Propene **IV** 2-methylbutene

b Edward wishes to produce a polymer that can be remoulded and recycled.

i What is the name of this type of polymer? (1 MARK)

ii Explain the difference between the polymer named in q19bi and elastomers, and explain why elastomers cannot be used for Edward's purpose. (3 MARKS)

Key science skills

Question 20 (4 MARKS)

Elisa wanted to test the properties of different types of plastics. As part of her experiment, she collected different items and tested the 'bendability' of the plastic. The results are provided.

Sample A	Sample B	Sample C
Was able to continually bend the plastic and it didn't rip or break.	Was very hard to bend the plastic. The plastic eventually snapped when I applied too much pressure to bend it.	Was able to bend the plastic backwards and forwards without it breaking.

a Identify the dependent variable. (1 MARK)

b What type of data is being collected in the experiment? (1 MARK)

c Comment on the reproducibility of the experiment. (2 MARKS)

Exam-style question hints

14: Square brackets can be used to indicate a repeating structure. **15:** The presence of branching impacts the density of a polymer. **16:** Heating thermoset polymers affects inter- and intramolecular bonds. **17:** Elastomers have elastic properties. **18a:** This reaction involves adding monomers together. **18b:** Pressure affects the stability of a plastic. **18c:** Decomposition involves the breaking down of intramolecular bonds. **19a:** Polymers are composed of a certain class of monomer. **19b:** Consider the plasticity of the desired polymer. **20a:** The dependent variable is affected by the independent variable. **20b:** The type of data is affected by the nature of the data collected. **20c:** Reproducibility measures the closeness of results between experiments in a certain condition.

6H MODIFYING POLYMERS

In this lesson, we will explore how polymers can be modified to carry out particular purposes.

6A Carbon materials | 6B Introduction to hydrocarbons | 6C Naming simple hydrocarbons | 6D Naming complex hydrocarbons | 6E Functional groups - alcohols and carboxylic acids | 6F Functional groups - esters | 6G Polymers | 6H Modifying polymers

Study design dot points

- the features of linear polymers designed for a particular purpose including the selection of a suitable monomer (structure and properties), chain length, degree of branching, percentage crystalline areas and addition of plasticisers
- the advantages and disadvantages of the use of polymer materials

Key knowledge units

Properties of polymers	1.2.12.1
Advantages and disadvantages of polymers	1.2.13.1

Key terms and definitions

Conductive polymers polymers able to conduct electricity

Copolymer polymers composed of two or more types of monomers

Crystalline region region of polymer with chains tightly packed together

Amorphous region region of polymer with random entanglement of chains

Isotactic all side chains are on one side of the polymer

Atactic side chains are randomly distributed along the polymer

Syndiotactic side chains on alternating sides of the polymer

Lesson links

This lesson builds on:

▶ 6A Carbon materials

Polymers are carbon materials.

▶ 6G Polymers

The structures discussed are polymers.

Properties of polymers 1.2.12.1

OVERVIEW

Polymers vary significantly in structure, and therefore possess different properties.

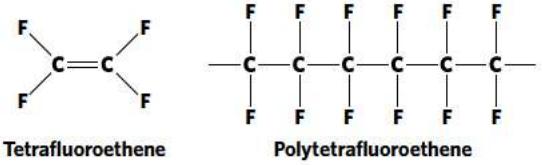
THEORY DETAILS

In the last lesson, we discussed the formation of polymers, as well as the structures and properties of particular types of polymers, with reference to their role in society. In this lesson we will be expanding upon these learnings by exploring modifications that can be made to polymers, and how these changes can enhance their utility.

Non-polar monomers

Since polymers are made up of many monomers bonded together in a sequence, the chemical properties of the monomers will affect those of the polymer. Therefore, the chemical characteristics of the monomers need to be taken into account to produce a polymer with particular properties.

Non-polar monomers are symmetrical which limits the monomers' ability to participate in dipole-dipole interactions. An example of a non-polar monomer is tetrafluoroethene, which is an ethene molecule with all the hydrogen atoms replaced by fluorine. The polymer formed from tetrafluoroethene is polytetrafluoroethene, also known as Teflon. The structures of tetrafluoroethene and polytetrafluoroethene are shown in figure 1.



Tetrafluoroethene **Polytetrafluoroethene**

Figure 1 Tetrafluoroethene and polytetrafluoroethene.

Since fluorine atoms are highly electronegative, polytetrafluoroethene exhibits unique properties such as being non-stick. The non stick property is due to the high electronegativity of fluorine which reduces bonding with other substances. One of the most common uses of Teflon is in frying pans, where the non-stick property is greatly beneficial. Teflon is also heat resistant which is useful in frying pans which are often subjected to high temperatures, as shown in figure 2.



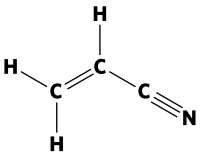
Image: preecha2531/Shutterstock.com

Figure 2 Teflon frying pan.

Polar monomers

If the monomers contain polar bonds and are asymmetrical, they are polar. When a polymer is formed from polar monomers, strong intermolecular bonds such as dipole-dipole interactions or hydrogen bonds form. As a result of the bonding, polar polymers are harder and more rigid than non-polar polymers.

Polyacrylonitrile is an example of a polymer composed of polar monomers. This polymer is used to make materials such as carbon fibre which is found in objects such as tennis racquets and fishing rods. Acrylonitrile monomers have one hydrogen atom replaced with a carbon atom triple bonded to nitrogen (C≡N). This carbon-nitrogen bond allows for strong dipole-dipole interactions to form between chains enabling materials such as carbon fibre to have high tensile strength, significant stiffness and low weight. This is shown in figure 3.



Acrylonitrile



Polyacrylonitrile material

Image: TLaoPhotography/Shutterstock.com

Figure 3 Acrylonitrile and polyacrylonitrile.

Conductive polymers

Conductive polymers are distinct from other polymers as they possess the capacity to conduct electricity. This property has enabled such polymers to be used to make polymer circuits, which are now being used in biosensors.

The polymer produced from the monomer ethyne, polyethyne, is able to conduct electricity under specific conditions. This is because the presence of alternating double and single bonds produces delocalised electrons, allowing a flow of electrons to produce an electrical current. The structure of polyethyne is shown in figure 4.

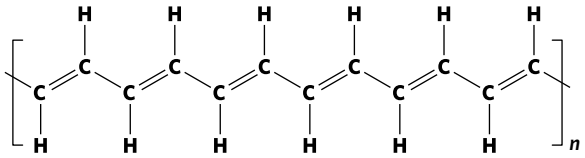
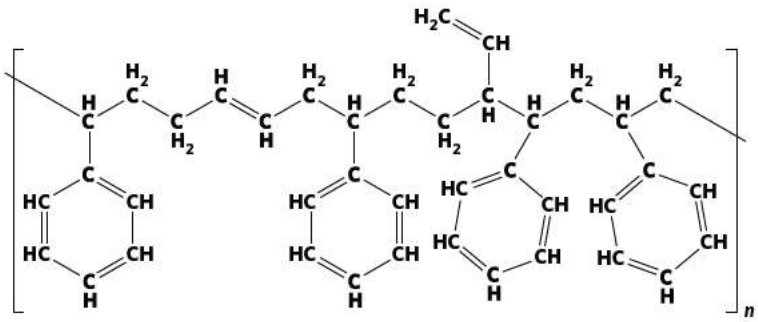


Figure 4 Polyethyne chain.

Copolymers

Copolymers are polymers that are composed of two or more different types of monomers. Since each type of monomer has distinct properties, altering the type and quantity of monomers in a polymer changes the properties of the entire molecule.

The polymer used to produce tyres is styrene-butadiene rubber (SBR), produced from the monomers styrene ($C_6H_5CH{=}CH_2$) and butadiene ($(CH_2{=}CH)_2$). The styrene/butadiene ratio of the polymer affects the properties of the tyre whereby increasing the amount of styrene increases the hardness and decreases how rubbery the material is. This property is useful in the context of car racing. A low styrene content produces soft tyres with strong grip. On the other hand, a high styrene content produces hard tyres enhancing durability, as shown in figure 5.



SBR chain



SBR-derived material

Image: Abdul Razak Latif/Shutterstock.com

Figure 5 SBR.

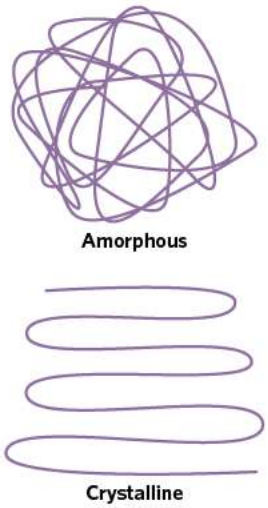
Crystallinity

Crystallinity refers to the structural regularity and orderliness of a molecule. The extent of crystallinity depends on the quantity of **crystalline regions** and **amorphous regions** in the polymer. Light interacts differently with crystalline and amorphous regions, providing these regions with distinct appearances.

Figure 6 depicts amorphous and crystalline polymers. Areas of a polymer where chains are arranged in a regular and orderly manner are known as crystalline regions. This uniform arrangement increases intermolecular bonding between chains and the ability for chains to pack closely together enhancing the strength and hardness of the substance. The strong bonds between chains means that minimal light is able to pass through the structure. As a result, crystalline regions appear opaque.

Amorphous regions occur where there is a random entanglement of polymer chains due to non-uniform and irregular arrangement of side chains. Consequently, polymer chains are unable to pack together tightly in amorphous regions. Polymers with many amorphous regions tend to be weaker, more flexible, and allow light to pass through. As a result, these regions appear translucent or transparent, rather than opaque.

Polymers generally consist of both crystalline and amorphous regions. By increasing the proportion of the polymer that is crystalline, the material becomes stronger, less flexible, and less transparent.



Amorphous

Crystalline

Figure 6 Amorphous and crystalline polymers.

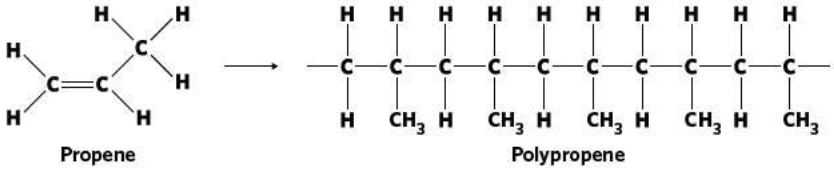
Effect of polymer chain length and branching

There can be significant variability in the polymer chain length and the degree of branching. The greater the chain length, the stronger the intermolecular bonds and toughness of the material. For example, ultra-high molecular weight polyethylene (UHMWPE) which has very large polymer chains and consequently strong dispersion forces is used to make bulletproof vests and hydraulic seals.

As discussed in lesson 6G, increasing the pressure during polymer production increases the degree of branching as is the case for low density polyethene (LDPE). On the other hand, low pressure during production which occurs for high density polyethene (HDPE) decreases branching. Hence, a high pressure environment resulting in extensive branching prevents the formation of crystalline regions meaning chains cannot pack tightly together. Consequently, the polymer is softer, more flexible and useful for products such as cling wrap.

Effect of side groups

Polypropene is a polymer made from the monomer propene, as shown in figure 7.

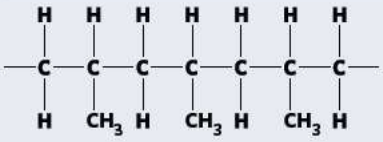
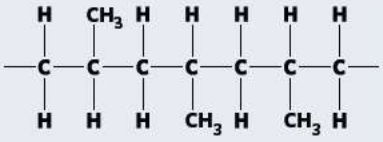
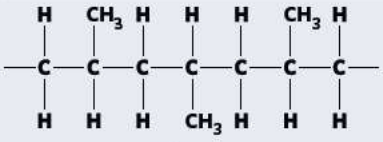


Propene Polypropene

Figure 7 Polypropene.

Although propene has a parent chain of three carbon atoms, only two of these carbon atoms will contribute to the carbon chain. Consequently, every second carbon atom will have a methyl side group in the polypropene polymer. The presence and orientation of the side chain affects the properties of the molecule. There are three forms of polypropene, which are described in table 1.

Table 1 Summary of the different types of polypropene.

Type of polymer	Description	Structure
Isotactic polypropene	All methyl side groups are found on one side of the polymer chain, which enables the formation of crystalline regions. This is a strong substance with a high melting point.	 Isotactic
Atactic polypropene	Methyl side groups are randomly distributed, preventing chains from bonding strongly together to form crystalline regions. Therefore, atactic polypropene is a soft substance with a relatively low melting point.	 Atactic
Syndiotactic polypropene	Methyl side groups are found on alternating sides of the polymer chain in a regular pattern. This enables the chains to pack together effectively, resulting in crystalline regions. Hence, this form of polypropene is a relatively hard substance with a high melting point.	 Syndiotactic

Size of side groups

Polymers that possess large side groups are unable to easily slide over each other or pack together, resulting in predominantly amorphous polymers. Polystyrene, for example, possesses a benzene ring (C_6H_5) on every second carbon atom in the polymer forming a hard, brittle and translucent substance, used for food containers and CD cases.

Additives

The majority of polymers include additives to enhance their properties, which include pigments for colour, stabilisers to prevent deterioration in sunlight, flame retardants and plasticisers.

Plasticisers are small molecules added during the production of polymers. These molecules reduce the density of the polymers by preventing chains from packing together tightly as shown in figure 8. Due to the weakened intermolecular forces, the material is softer and more flexible. For example, the addition of plasticiser to PVC, which is usually hard and rigid due to the strong dipole-dipole interactions, enables the chains to slide past each other, increasing the softness and flexibility of the substance.

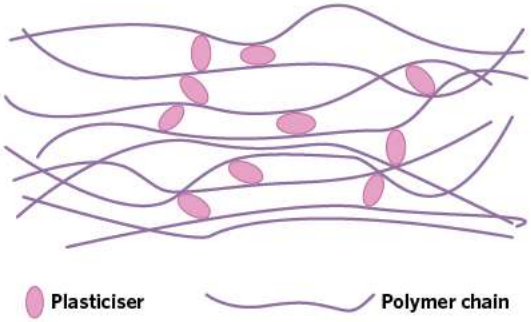


Figure 8 The function of plasticisers.

Foamed polymers are produced by blowing gas through melted polymers as shown in figure 9. For example, polystyrene foam is created by passing pentane gas through polystyrene beads when in a liquid state. The gas that is introduced makes the polymers swell to produce a foam. Due to their lightweight, shock-absorbing and insulating properties, polymer foams are often used for packaging materials.



Image: Anna Hoychuk/Shutterstock.com

Figure 9 Polystyrene foam.

Advantages and disadvantages of polymers 1.2.13.1

OVERVIEW

While polymers are exceptional in their usefulness and versatility, their creation and use cause harm to the environment.

THEORY DETAILS

Due to the wide range of applications of polymers, they are commonly used. However, we also need to be aware of the disadvantages of using these substances. The main advantages and disadvantages of polymers are shown in table 2.

Table 2 Advantages and disadvantages of polymers.

Advantages	Disadvantages
Able to be produced from renewable sources such as sugarcane.	Primarily produced from fossil fuels such as petroleum.
Highly modifiable and diverse in characteristics such as colour.	Unable to be broken down by microorganisms and can pollute the environment.
Cheap to produce.	Some polymers are easy to crack or break.
Low density, lightweight, strong and chemically resistant.	May release toxic fumes when burned.
Some plastics are recyclable.	Not all plastics are recyclable.

Australians use approximately 1.5 billion kg worth of polymer-based materials annually. Since many polymers are unable to be biodegraded and are resistant to weathering, plastics can remain in the environment for a very long time. Consequently, plastics occupy landfills and pollute oceans and waterways which damages ecosystems. Due to the toxic fumes produced when plastics are burnt, this is not an environmentally friendly method of removal from the environment.

Biodegradable plastics

Polymers that can be broken down naturally by microorganisms are referred to as biodegradable plastics. These are a new type of polymer that are not yet widely used. However, increased awareness of the damaging effects of polymers on the environment has increased the demand of these biodegradable forms. Most biodegradable plastics are produced from renewable plant materials such as corn or starch.

Another method for producing biodegradable polymers is through the incorporation of additives into polymers, which promote the degradation of polymer chains into small compounds over time.

Polymer recycling

The problems associated with the accumulation of polymers in landfill and poor biodegradation has led to the prioritisation of polymer recycling. The primary methods of recycling polymers are:

- Reprocessing – Involves shredding, melting and reshaping plastic into new products. This is only possible for thermoplastic polymers.
- Depolymerisation – Polymers are broken down into monomers, which are used to produce new polymers. This process requires significant amounts of energy and produces a low yield, and is therefore less economically viable than reprocessing.

Plastic materials display a particular symbol to signify whether and how the material can be recycled, as shown in figure 10.

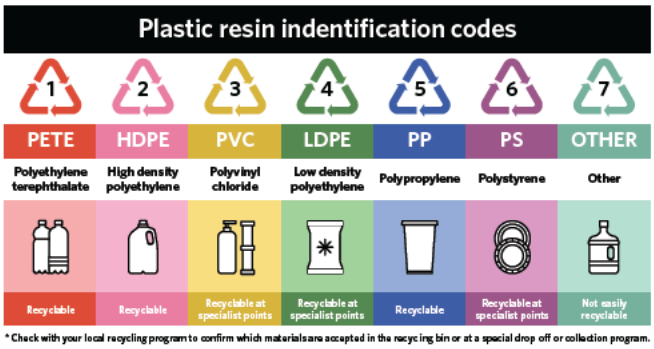


Image: shopplaywood/Shutterstock.com

Figure 10 The codes used to identify plastics that can be recycled.

Theory summary

- Monomers determine the properties of polymers such as non-stick, hardness and electrical conductivity.
- Copolymers are produced from two or more different types of monomers.
- Polymer chain length, degree of branching and size of side groups all impact the intermolecular interactions of polymers.
- Plasticisers can be used to make materials softer and more flexible.
- Despite the advantages of polymers, there are significant environmental drawbacks.
- Some polymers can be biodegraded or recycled, but these methods have limitations.

6H QUESTIONS

Theory review questions

Question 1

Copolymers

A consist of multiple polymers bound together.

B are polymers containing more than one type of monomer.

Question 2

Teflon is a non-polar polymer used

A in pipes due to its hardness.

B on pans due to its non-stick property.

Question 3

Crystallinity is

A when polymers pack together closely and become opaque.

B only found in polymers with strong dipole-dipole interactions.

Question 4

Which of the following describes the three types of polypropene accurately?

A

Isotactic	Atactic	Syndiotactic
Side chains regularly alternating	All side chains on the same side	Side chains random

B

Isotactic	Atactic	Syndiotactic
All side chains on the same side	Side chains random	Side chains regularly alternating

Question 5

Plasticisers

A soften polymers by weakening intermolecular forces.

B assist polymers to harden into solid plastics.

Question 6

Which of the following is true regarding polymers?

A Most polymers are produced from renewable sources.

B It is possible to recycle some forms of polymers.

Question 7

The types of polymer recycling are

A reprocessing and depolymerisation.

B repolymerisation and degradation.

Question 8

Increasing the chain length of polymers

A increases strength due to increased intermolecular bonding.

B decreases strength due to increased packing.

Question 9

Increasing branching in polymers

A decreases the number of crystalline regions.

B increases intermolecular bonding.

Question 10

Polar monomers

A repel non-polar polymers.

B increase the hardness of polymers.

Deconstructed questions

Use the following information to answer questions 11–13.

Gustav is a chemistry student who has been working on developing a new type of polymer. He asks you some questions about how to design the polymer.

Question 11 (1 MARK)

What would be the effect of using polar monomers as opposed to non-polar monomers to produce a polymer?

A The material will be more flexible as strong bonds will be maintained despite bending.

B The material will be weaker as polymers will repel each other.

C The material will be stronger due to stronger intermolecular bonding.

D The material will be more flammable as asymmetrical molecules have a greater tendency for burning.

Question 12 (1 MARK)

Why are plasticisers effective in making polymers softer?

A Plasticisers lead to weaker intermolecular forces.

B Plasticisers lead to stronger intermolecular forces.

C Plasticisers prevent polymer degradation.

D Plasticisers produce more crystalline regions.

Question 13 (4 MARKS)

Gustav is interested in the consequences of using polymer materials. Identify two advantages and two disadvantages of polymers.

Exam-style questions

Within lesson

Question 14 (1 MARK)

Increasing the size of side groups in monomers

A enhances the ability of a polymer to be recycled.

B increases the flexibility of a material.

C produces a softer material.

D prevents polymer chains from sliding over each other.

Question 15 (8 MARKS)

Wolfgang is a scientist who is experimenting with different types of polymers. He wishes to determine the effects of increasing polymer chain length and branching.

a Wolfgang developed two polymers with the following structure. Compare the properties of each. (2 MARKS)

Polymer 1  **Polymer 2** 

b After changing some of the steps in his methodology, Wolfgang was able to produce the following polymers. Compare the properties of each. (2 MARKS)

Polymer 1  **Polymer 2** 

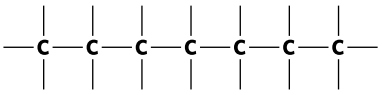
c Wolfgang decides to produce a polymer with a small amount of branching and long polymer chains. Identify a potential use of this polymer, justifying your answer. (4 MARKS)

Question 16 (8 MARKS)

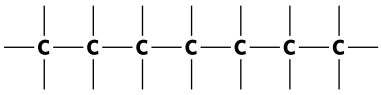
Ludwig is using polypropene to produce a polymer, and is investigating how this type of polymer is unique from other polymers.

a What is the primary characteristic of polypropene that makes it unique from polyethene? (1 MARK)

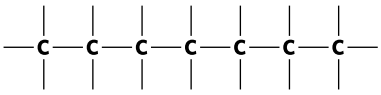
b Fill in the side chains for each of the following molecules: (3 MARKS)



Isotactic polypropene



Atactic polypropene



Syndiotactic polypropene

c Ludwig produces a soft and flexible material made from a single type of polypropene.

i Which type of polypropene is this most likely to be? Explain. (2 MARKS)

ii Can assessing the softness and flexibility of a material be used to accurately identify what polymer it is made from? Explain. (2 MARKS)

Multiple lessons

Question 17 (9 MARKS)

Polymers are highly versatile molecules that are commonly used. However, these materials do have their drawbacks.

a What is the primary source of polymers? Identify at least one other source of polymers, and compare their renewability. (3 MARKS)

b Some polymers are biodegradable.

i What does the term 'biodegradable' mean with reference to polymers? (1 MARK)

ii Identify at least one method of making polymers biodegradable. (1 MARK)

c Another method of reducing the impact of polymers on the environment is through recycling.

i Describe the two primary methods of recycling polymers. (2 MARKS)

ii Explain the primary limitations of each method of recycling. (2 MARKS)

Question 18 (13 MARKS)

Claude is investigating how polymers can be modified to possess different characteristics, and decides to explore various methods of polymer modification.

a What does the term 'copolymers' refer to? Describe their structure and how they can be altered. (3 MARKS)

b Claude wishes to produce a copolymer that is conductive. In carrying out his research, Claude compares the conductivity of polyethene and polyethyne.

i What is a conductive polymer? (1 MARK)

ii Draw the structure of the monomer of polyethyne. (1 MARK)

iii Determine whether polyethene and polyethyne are saturated or unsaturated molecules. Briefly explain your answer. (2 MARKS)

iv Briefly describe why polyethyne is a conductive polymer while polyethene is not. (2 MARKS)

c In continuing his studies of polymer modifications, Claude considers the use of additives to further alter the properties of the molecule, for example plasticisers.

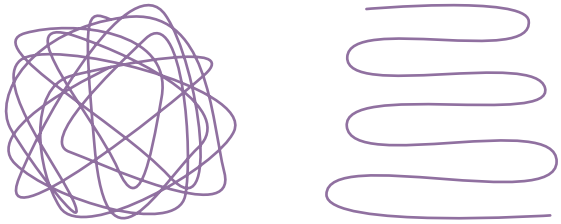
i Describe how the addition of plasticisers affects the interactions between polymer chains. (2 MARKS)

ii Identify how the properties of a polymer with plasticisers will differ from those of a polymer without plasticisers. (2 MARKS)

Question 19 (12 MARKS)

A student wishes to explore the concept of crystallinity in polymers. In order to observe the differences in properties between crystalline and amorphous polymer regions, the student produces a series of polymers under different conditions.

a Label the following structures as either crystalline or amorphous. (2 MARKS)



b Identify whether LDPE or HDPE is more likely to form crystalline or amorphous regions. (2 MARKS)

c Explain the conditions under which LDPE and HDPE are produced, and how this is reflected in their structures. (3 MARKS)

d Explain how LDPE and HDPE interact with light. (3 MARKS)

e The student wishes to produce a flexible material from a polymer. Identify whether it would be best to produce LDPE or HDPE. Explain your answer. (2 MARKS)

Key science skills

Question 20 (7 MARKS)

Franz is a chemical engineer designing a system for producing different types of polymers. He created polymers using each of the techniques and compares the characteristics of each type of polymer. Franz tested each process four times under identical conditions and presented his results in the table shown.

	Process 1	**Process 2**	**Process 3**
Trial 1	Hard, brittle	Soft, flexible	Hard, brittle
Trial 2	Soft, flexible	Soft, flexible	Hard, brittle
Trial 3	Soft, flexible	Soft, flexible	Medium, brittle
Trial 4	Soft, brittle	Soft, flexible	Hard, brittle

a Define repeatability and comment on the repeatability of each of the processes. (4 MARKS)

b Franz knows that the only difference between the processes is the amount of pressure that each is exposed to. Determine which process is at the highest and which is at the lowest pressure. Justify your answer. (3 MARKS)

Exam-style question hints

14: Side groups occupy space and interact with each other. **15a:** Increasing chain length increases the strength of intermolecular bonding. **15b:** Branching reduces the linearity of polymers and increases the space occupied by the molecule. **15c:** Various factors influence bonding strength. **16a:** Propene has a parent chain with three carbon atoms. **16b:** Side chains in different polypropene structures are arranged differently. **16c:** The positioning of side groups affects the properties of a polymer. **17a:** Polymers are often non-renewable. **17b:** The 'bio' in biodegradation refers to biological activity. **17c:** Recycling involves breaking down an old material to form a new material. **18a:** 'Co-' means 'joint'. **18bi:** Conductive polymers require free electrons. **18bii:** The suffix of a hydrocarbon indicates the type of bonding between carbons in the carbon chain. **18biii:** Saturation refers to the types of carbon-carbon bonds present in a hydrocarbon. **18biv:** Some polymers conduct electricity depending on the bonds present. **18ci:** Additives are present in between polymer chains. **18cii:** Plasticisers disrupt normal bonding processes. **19a:** Crystalline and amorphous describe different structures. **19b:** Crystalline structures prevent the transmission of light through the polymer. **19c:** High pressures prevent the formation of uniform and regular polymer chains. **19d:** The density of polymers affects interactions with light. **19e:** Flexible materials have relatively weak intermolecular bonding. **20a:** Reliability encompasses both repeatability and reproducibility. **20b:** High pressures prevent the formation of neat polymer chains.

EXPERIMENT

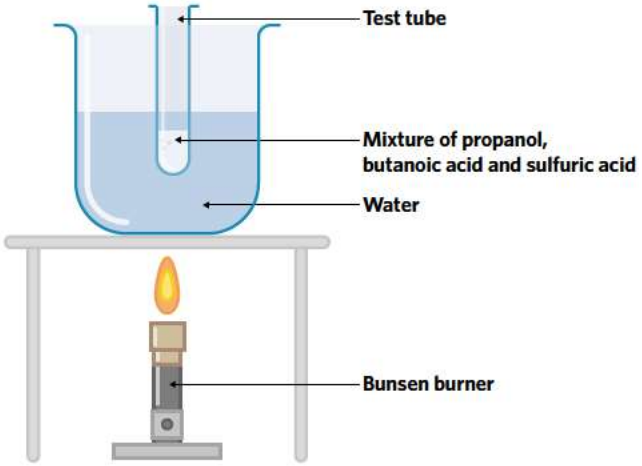
PREPARATION OF AN ESTER

In this chapter, we have investigated the properties and reactivity of different hydrocarbons. Hydrocarbons are organic molecules which vary by the length of their carbon chains and the presence or absence of functional groups.
The experiment involves reacting and comparing hydrocarbons with different functional groups.

Materials

(per group)

- 2 × 250 mL glass beakers
- 5 × Test tubes
- 1 × Bunsen burner
- 1 × Thermometer
- 10 mL of 1.0 M ethanol, CH_3CH_2OH
- 10 mL of 1.0 M propan-1-ol, $CH_3CH_2CH_2OH$
- 10 mL of 1.0 M butan-1-ol, $CH_3CH_2CH_2CH_2OH$
- 10 mL of 1.0 M propanoic acid, CH_3CH_2COOH
- 10 mL of 1.0 M butanoic acid, $CH_3CH_2CH_2COOH$
- Ice cubes
- Concentrated sulfuric acid, H_2SO_4



Note: Depending on the concentration of the sulfuric acid, it is most likely that your teacher will be the only person who is able to handle this reagent.

Method

1. Add 125 mL of water to a 250 mL beaker, and heat the water to 60°C using a bunsen burner.
2. Add 10 mL of an alcohol (ethanol, propan-1-ol or butan-1-ol) and 10 mL of a carboxylic acid (ethanoic acid, propanoic acid, butanoic acid) and 3 drops of concentrated sulfuric acid to a new test tube.
3. Place the test tube in the heated water for 15 minutes. Ensure that the water remains at approximately 60°C throughout.
4. While waiting, add 150 mL of tap water into another 250 mL beaker, along with some cubes of ice.
5. Once the test tube has been left for 15 minutes, transfer the test tube from the heated water to the beaker containing the ice.
6. Add 5mL of distilled water to the test tube. Note the odour produced in the experiment by wafting the fumes towards your nose.
7. Repeat steps 2–6 with different combinations of alcohols and carboxylic acids and record any observations.

Results

Reactants	Observations

QUESTIONS

Question 1 (3 MARKS)

Complete the following table.

Alcohol	Carboxylic acid	Product
Ethanol	Propanoic acid	
		Propyl butanoate

Question 2 (2 MARKS)

Explain why water is also produced in the reaction.

Question 3 (2 MARKS)

Explain the role of sulfuric acid in this experiment.

Question 4 (2 MARKS)

Describe what would take place if the test tube were not added to heated water.

ANSWERS

1

Alcohol	Carboxylic acid	Product
Ethanol	Propanoic acid	Ethyl propanoate
Propan-1-ol	Butanoic acid	Propyl butanoate

2 [Water is produced as a byproduct in this reaction as this is a condensation reaction.[1]][In condensation, multiple reactants react together to form a single larger product, with a OH from a carboxyl group and a H from a hydroxyl group being used to form water.[2]]

I have identified that this is a condensation reaction.[1]

I have explained what takes place in a condensation reaction.[2]

3 [Sulfuric acid acts as a catalyst in the reaction between propanol and butanoic acid.[1]][This means that the presence of sulfuric acid provides a lower energy pathway for the reaction to take place without expending as much energy.[2]]

I have identified that sulfuric acid works as a catalyst.[1]

I have explained the role of a catalyst in the experiment.[2]

4 [Without introducing the test tube to heated water, there would be insufficient energy to efficiently carry out the reaction.[1]][Because of this, the reaction would either take place at a very slow rate or would not take place at all.[2]]

I have identified what would happen if the test tube was not added to heated water.[1]

I have explained why this would happen.[2]

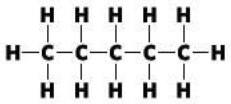
CHAPTER 6 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Question 1 (1 MARK)

Which of the following shows the structure of diamond?

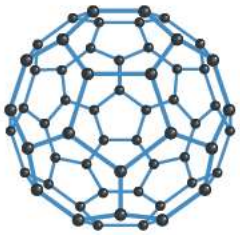
A



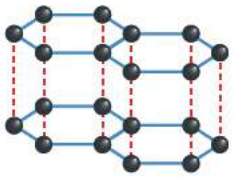
B



C



D



Question 2 (1 MARK)

Graphene is a nanomaterial that

A conducts electricity due to delocalised electrons.

B weak as it consists of only a single layer of atoms.

C is shaped like a hollow sphere.

D is inflexible as it consists of multiple layers.

Question 3 (1 MARK)

Crude oil is

A a renewable source of hydrocarbons.

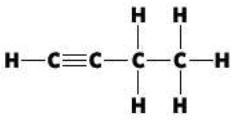
B hydrocarbons obtained from the fractional distillation of plant oils.

C hydrocarbons obtained from the earth using an oil well.

D a non-hydrocarbon form of oil.

Question 4 (1 MARK)

The molecule shown is considered as



A an isomer.

B an alkene.

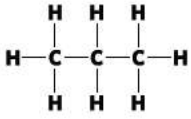
C an alkyne.

D a dimer.

Question 5 (1 MARK)

Which of the following is not a representation of propane:

A



B $CH_3CH_2CH_3$

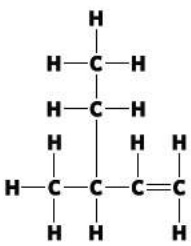
C



D C_3H_{10}

Question 6 (1 MARK)

The name of the following molecule is



A 3-methylpent-1-ene.

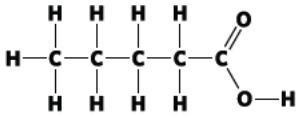
B 3-ethylbut-1-ene.

C 2-ethylbut-3-ene.

D 3-ethylbut-1-yne.

Question 7 (1 MARK)

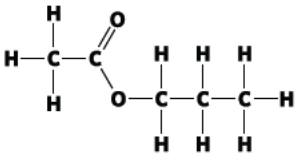
The name of the following molecule is



A pentanol.
B pentanoic acid.
C pentan-1-ol.
D pentan-1-oic acid.

Question 8 (1 MARK)

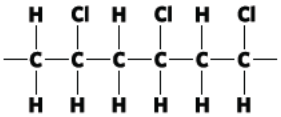
The name of the following molecule is



A propyl ethanoate.
B butyl methanoate.
C ethyl propanoate.
D methyl butanoate.

Question 9 (1 MARK)

The following is a section of polymer.



The name of the monomer of this polymer is

A chloroethane.
B 2-chloroethane.
C chloroethene.
D chloro-2-methane.

Question 10 (1 MARK)

Crystalline regions of polymer materials are formed

A in polymers with extensive branching.
B in polymers with strong intermolecular bonding.
C in polymers which enable the transmission of light.
D in polymers with long carbon chains.

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (13 MARKS)

Crystal is a year 11 chemistry student who wants to study the interactions and properties of hydrocarbons with different functional groups. To begin with, she collects samples of ethanol and propanoic acid.

a Draw the structural formulas of ethanol and propanoic acid. (2 MARKS)

b Crystal produces samples of pure liquid ethanol and propanoic acid and places them in test tubes. Following this, she increases the temperature of each sample, and records the temperature at which each begins boiling.

i Which sample would be expected to have a higher boiling point? Explain your answer. (2 MARKS)

ii Describe how the boiling points of ethanol and propanoic acid compare to ethane. (2 MARKS)

c Crystal wishes to investigate the reactivity of ethanol and propanoic acid. She produces two beakers containing ethanol and propanoic acid under identical conditions, however an extra compound is added to one of the beakers. Crystal notes that a fruity-smelling substance begins to form in one of the beakers, but not in the other.

i Identify the chemical name of the main substance produced. (1 MARK)

ii Draw the structural formula of the main substance produced, circling the functional group of the molecule. (1 MARK)

iii Identify the substance that was added to the beaker where the reaction was successful. (1 MARK)

iv What is the specific and the general name of the reaction that has taken place between propanoic acid and ethanol? (1 MARK)

d Crystal collects the product of the reaction between propanoic acid and ethanol and separates it into two beakers – one that is exposed to heat to determine the boiling point, and another where water is introduced.

i Identify whether the boiling point of the product would be higher or lower than propanoic acid and ethanol. (1 MARK)

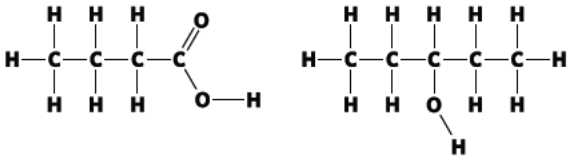
ii Describe what would be expected to take place when water is introduced at a high temperature in the presence of an acid. (2 MARKS)

Question 12 (8 MARKS)

Robert is a marine scientist observing the impact of using hydrocarbon products, most specifically polymers, on marine ecology. To better understand the properties of these materials, he decides to study the origins and manufacturing processes of each.

a Robert has identified a substantial number of hydrocarbons in ocean waters. He understands that many of these may have been introduced from crude oil spills, as well as production of hydrocarbons by cyanobacteria. To perform further analysis, he collects a sample of hydrocarbon molecules.

i Provide the names of the molecules shown. (2 MARKS)



ii Robert notices that one of the molecules identified in q12ai is a positional isomer. Positional isomers are isomers that differ only in the location of their functional group. Draw and name the other isomers of this molecule. (4 MARKS)

b Robert has obtained a sample of polymers from the ocean, and wants to investigate recycling.

Robert's first sample is a thermoplastic polymer. Explain the possible recycling processes for this polymer. (2 MARKS)

Question 13 (9 MARKS)

Linda is interested in the different types of carbon materials and nanomaterials that can be produced.

a Linda begins her observations by comparing the structures and properties of diamond and graphite.

i Describe the structures of diamond and graphite. (2 MARKS)

ii Which of these materials is most likely to be used in products that conduct electricity? Justify your answer. (2 MARKS)

b Linda is interested in comparing graphite and graphene.

i Identify the primary difference between graphite and graphene. (1 MARK)

ii Graphene is often described as a 'wonder material'. Provide some examples of why graphene is described this way. (1 MARK)

c Linda understands that another prominent carbon nanomaterials is fullerenes.

Identify two commonly used carbon nanomaterials, describing their shapes and potential uses. (3 MARKS)

KEY SCIENCE SKILLS QUESTIONS (4 MARKS)

Question 14 (4 MARKS)

Ed is a science writer for a reputable science magazine. He is designing an experiment that could be carried out to test the effectiveness of industrial equipment. Ed's experiment involves the formation of diamonds from carbon under extreme pressures and temperatures. Ed sends the experiment out to a number of manufacturers to be tested.

a Three manufacturers with identical equipment and correctly following Ed's method had three wildly different results, with diamonds only being formed in one case.

Is this experiment a reliable measure of the effectiveness of industrial equipment? Justify your answer. (2 MARKS)

b Frustrated by criticisms of his experimental technique, Ed sets out to determine why the trials indicated that his experiment was unreliable. Ed soon discovers that the technique used by the manufacturing companies to determine whether diamonds had formed were different.

Explain why this may be a problem. Provide an example. (2 MARKS)

UNIT

2 What makes water such a unique chemical?

Water is the most widely used solvent on Earth. In this unit students explore the physical and chemical properties of water, the reactions that occur in water and various methods of water analysis.

Students examine the polar nature of a water molecule and the intermolecular forces between water molecules. They explore the relationship between these bonding forces and the physical and chemical properties of water. In this context students investigate solubility, concentration, pH and reactions in water including precipitation, acid-base and redox.

Students are introduced to stoichiometry and to analytical techniques and instrumental procedures, and apply these to determine concentrations of different species in water samples, including chemical contaminants. They use chemistry terminology including symbols, units, formulas and equations to represent and explain observations and data from experiments, and to discuss chemical phenomena. Students explore the solvent properties of water in a variety of contexts and analyse selected issues associated with substances dissolved in water.

A practical investigation into an aspect of water quality is undertaken in Area of Study 3. The investigation draws on content from Area of Study 1 and/or Area of Study 2.

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UNIT 2

AOS 1

How do substances interact with water?

In this area of study students focus on the properties of water and the reactions that take place in water including acid-base and redox reactions. Students relate the properties of water to the water molecule's structure, polarity and bonding. They also explore the significance of water's high specific heat capacity and latent heat of vaporisation for living systems and water supplies.

Students investigate issues associated with the solubility of substances in water. Precipitation, acid-base and redox reactions that occur in water are explored and represented by the writing of balanced equations.

Students compare acids with bases and learn to distinguish between acid strength and acid concentration. The pH scale is examined and students calculate the expected pH of strong acids and strong bases of known concentration.

Outcome 1

On completion of this unit the student should be able to relate the properties of water to its structure and bonding, and explain the importance of the properties and reactions of water in selected contexts.

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UNIT 2 AOS 1, CHAPTER 7

Water

07

7A Properties of water

7B Thermal properties of water

7C Dissolution

7D Concentration

7E Solubility

7F Precipitation reactions

Key knowledge

- trends in the melting and boiling points of Group 16 hydrides, with reference to the nature and relative strengths of their intermolecular forces and to account for the exceptional values for water
- specific heat capacity and latent heat including units and symbols, with reference to hydrogen bonding to account for the relatively high specific heat capacity of liquid water, and significance for organisms and water supplies of the relatively high latent heat of vaporisation of water
- the comparison of solution processes in water for molecular substances and ionic compounds
- precipitation reactions represented by balanced full and ionic equations, including states
- the importance of the solvent properties of water in selected biological, domestic or industrial contexts
- the use of solubility tables and experimental measurement of solubility in gram per 100 g of water
- the quantitative relationship between temperature and solubility of a given solid, liquid or gas in water
- the use of solubility curves as a quantitative and predictive tool in selected biological, domestic or industrial contexts
- the concept of solution concentration measured with reference to moles (mol L^{-1}) or with reference to mass or volume (g L^{-1}, mg L^{-1}, %(m/m), %(m/v), %(v/v), ppm, ppb) in selected domestic, environmental, commercial or industrial applications, including unit conversions

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7A PROPERTIES OF WATER

Water is all around us and is essential for life. The properties of water can explain why water is such a critical chemical compound and why we are here today.

7A Properties of water | 7B Thermal properties of water | 7C Dissolution | 7D Concentration | 7E Solubility | 7F Precipitation reactions

Study design dot points

- trends in the melting and boiling points of Group 16 hydrides, with reference to the nature and relative strengths of their intermolecular forces and to account for the exceptional values for water
- the comparison of solution processes in water for molecular substances and ionic compounds
- the importance of the solvent properties of water in selected biological, domestic or industrial contexts

Key knowledge units

Properties of Group 16 hydrides	2.1.1.1
Application of water as a solvent	2.1.5.1
Significance of the properties of water	2.1.2.3

Key terms and definitions

Group 16 hydrides Group 16 element covalently bonded to hydrogen

Solution mixture where a solute has dissolved in a solvent

Aqueous solution solution in which the solvent is water

Solute substance that dissolves in a solvent

Solvent substance that dissolves another substance, usually water

Homogenous same proportion of particles throughout the solution (i.e. well-mixed)

Ionisation process by which a chemical species gains or loses an electric charge

Hydrated (ions) ions that become surrounded by water molecules

Ion-dipole attraction electrostatic attraction between a molecule's permanent dipole and an ion

Lesson links

This lesson builds on:

▶ 5C Intermolecular bonding

The properties of water are linked to its intermolecular bonding abilities.

Properties of Group 16 hydrides 2.1.1.1

OVERVIEW

When compared with Group 16 hydrides, water has unique and distinctive properties.

THEORY DETAILS

Water has the molecular formula H_2O and is a covalent compound. It consists of a central oxygen atom covalently bonded to two hydrogen atoms. As we learnt in lesson 5A, a covalent bond involves the sharing of electrons and so both hydrogen atoms share one electron with the central oxygen atom. This is demonstrated in the Lewis structure shown in figure 1.



Figure 1 Lewis structure of H_2O.

Furthermore, as we learnt in lesson 5B, pairs of negatively-charged electrons repel each other and will therefore adopt a configuration with each pair as far apart from each other as possible. In this case, since we have four pairs of electrons, the electron pairs will adopt a tetrahedral shape around the central oxygen atom. Since two pairs are lone pairs of electrons (electrons not involved in bonding), water molecules will be V-shaped or bent, as shown in figure 2.



Figure 2 Ball-and-stick model of water, with the oxygen atom in red and hydrogen atoms in white.

Since the oxygen atom is highly electronegative, it attracts the shared electrons in the covalent bonds with both hydrogen atoms towards itself, resulting in polar covalent bonds. As water molecules are asymmetrical in three-dimensions, each of the dipoles can be added together, resulting in the overall permanent dipole shown in figure 3. This means that water molecules are polar, which will influence many of their physical and chemical properties.

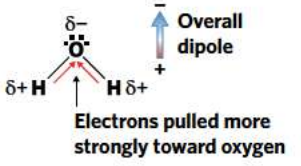


Figure 3 Overall permanent dipole in a water molecule.

Consequently, water is able to form permanent dipole-dipole bonds with itself since it is polar and has a permanent dipole. The strongest intermolecular forces present between water molecules are hydrogen bonds. Hydrogen bonding occurs between water molecules as the O-H bonds in water are highly polar which results in the oxygen atom having a partial negative charge, and the hydrogen atoms having a partial positive charge. When these partial charges interact, there is an electrostatic attraction.

In fact, since the oxygen atom has two lone pairs of electrons which are dense in negative charge, these two sites can hydrogen bond with the hydrogen atoms on neighbouring molecules. The two partially positively charged hydrogen atoms can also form hydrogen bonds with the oxygen atoms of neighbouring water molecules. Therefore, each water molecule can form up to four hydrogen bonds with four other water molecules, as shown in figure 4.

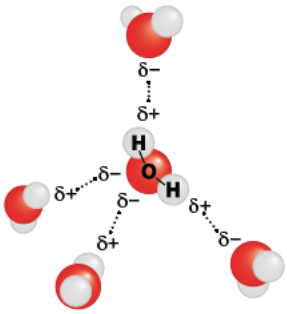


Figure 4 Water molecule forming hydrogen bonds with four other water molecules.

Group 16 hydrides

Group 16 hydrides are compounds that consist of a Group 16 element that is covalently bonded to two hydrogen atoms (it is usually two due to the octet rule, but exceptions exist like H_2O_2). Group 16 elements are shown in figure 5, and therefore possible Group 16 hydrides include water (H_2O), hydrogen sulfide (H_2S), hydrogen selenide (H_2Se), hydrogen telluride (H_2Te) and hydrogen polonide (H_2Po).



Figure 5 Elements in Group 16 of the periodic table.

Since Group 16 hydrides are all molecular compounds, the temperature of their melting and boiling points is representative of the strength of their intermolecular forces. That is, molecules with very strong intermolecular forces will have higher melting and boiling points. It would be expected that both melting and boiling points would increase as we move down Group 16. This is because the strength of the intermolecular dispersion forces increases with the size of the molecules. While this is partially true as shown by figure 6, water is a unique molecule in that it has the highest melting point and boiling point of all Group 16 hydrides.

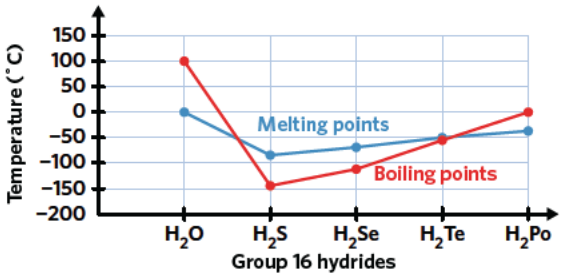


Figure 6 Graph of the melting and boiling points of Group 16 hydrides.

The same difference can be seen not just with other Group 16 hydrides but also other molecules of similar sizes, as shown in figure 7.

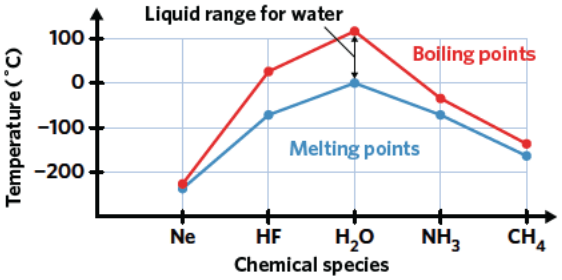


Figure 7 Graph of the melting and boiling points of other similarly sized molecules.

This unexpected result is due to water's unique ability to form hydrogen bonds with four other water molecules. These bonds require a significant amount of energy to disrupt.

Another interesting property of water that arises from its unique hydrogen bonding capabilities is that when the temperature of water reaches its freezing point at 0°C, the water molecules start to arrange themselves into a regular lattice, with each water molecule bonded to four other water molecules. This means that water molecules are spaced further apart from each other as a solid than they would be in the liquid state, as shown in figure 8.

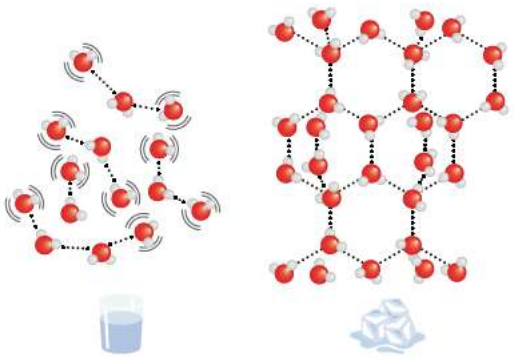


Figure 8 Arrangement of water molecules when in the liquid state (left) compared to ice, the solid state (right).

The arrangement of water molecules in a regular lattice results in ice being less dense and able to float on water. This was explored in detail in lesson 5C.

Application of water as a solvent 2.1.5.1

OVERVIEW

One of water's most important properties is that it is able to dissolve substances, which makes water a solvent.

THEORY DETAILS

A **solution** is a mixture of two or more different substances where one is dissolved in another. When water is used to dissolve a substance, which could be a solid, liquid, or gas, an **aqueous solution** is produced. In an aqueous solution, the particles are free to move about, mix and react. The chemical substance that is dissolved in the solution is known as the **solute** and the substance that it is dissolved in is the **solvent,** like water. Water is always the solvent in aqueous solutions, but there are other types of solutions that exist which use different solvents.

One of the defining characteristics of a solution is that it is **homogenous**. This means that the solute and solvent are mixed together and cannot be differentiated from each other. For example, sea water is a solution of salt (solute) and water (solvent) because they are evenly mixed. A mixture of sand and water, however, is not a solution because sand and water can be easily separated from each other. Other properties of solutions are that the dissolved particles cannot be seen and that different solutions can have different amounts of solute dissolved in them.

When a substance dissolves, there are three different kinds of intermolecular forces that need to be considered:

1 Intermolecular forces between solvent particles.
2 Intermolecular forces between solute particles.
3 Intermolecular forces between solute and solvent particles.

This is because in order for a substance to dissolve, the intermolecular forces between solute particles and between solvent particles need to be overcome by the intermolecular forces between the solute and solvent. If the solute-solvent forces are very weak, or the solute-solute and/or solvent-solvent forces are much stronger, a substance will not dissolve.



For a substance to dissolve, the solute-solute and solvent-solvent forces need to be overcome.

This process is illustrated in figure 9.

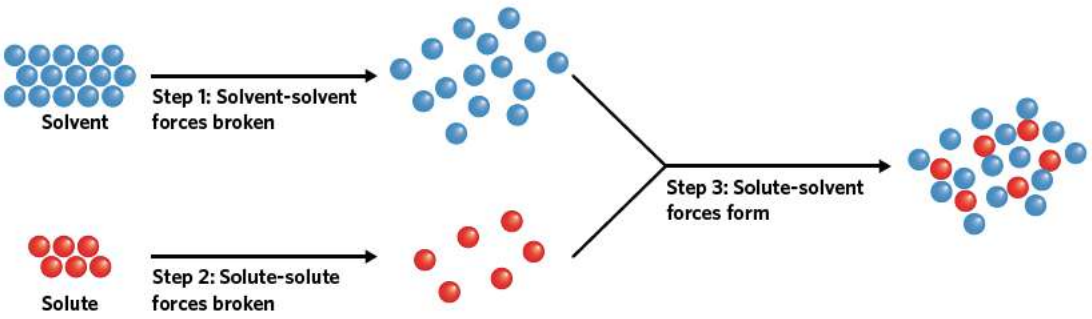


Figure 9 Solute-solute and solvent-solvent forces need to be overcome before a molecule dissolves.

The three different kinds of intermolecular forces will be discussed in greater detail later in this lesson but for now it is sufficient to know that as a general principle, 'like dissolves like'. This means that polar compounds like water will generally dissolve compounds which are polar or charged, but will not dissolve non-polar compounds. This explains why pentane (non-polar) and water (polar) are unable to mix but methanol (polar) can dissolve in water.

Most molecular substances are insoluble in water. This is because the solute-solvent forces between water and these molecules are insufficient to overcome the stronger intermolecular bonding present between the water molecules and between solute molecules. Despite this, there are some molecules that are soluble in water, due to two main mechanisms:

- The molecule forms hydrogen bonds with water molecules.
- The molecule ionises in water.



'Like dissolves like' is an important principle in determining whether a compound is able to dissolve in water or not.

Hydrogen bonding with water

Some molecules have a functional group that is capable of hydrogen bonding with water molecules (for example, the hydroxyl, carboxyl and ester functional groups). As we learnt in chapter 6, the influence of a functional group on a molecule decreases as the chain length increases. Therefore, only small molecules with these functional groups, or large molecules with many of these functional groups, tend to dissolve in water.

Glucose ($C_6H_{12}O_6$) is a highly polar molecule that is able to dissolve in water. The intermolecular forces that exist between glucose molecules are dispersion forces, permanent dipole-dipole forces and hydrogen bonds as shown in figure 10.

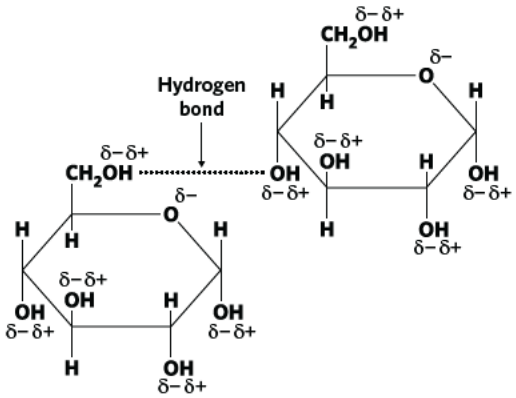


Figure 10 Hydrogen bonding between two glucose molecules, and other possible hydrogen bonding sites.

In order for glucose to dissolve, the intermolecular hydrogen bonds between glucose molecules need to be disrupted, as well as the intermolecular hydrogen bonds between water molecules, shown in figure 11. This is possible because the new interactions that form between glucose and water molecules are also hydrogen bonds and of a similar strength. Therefore, the solute-solute and solvent-solvent intermolecular forces are able to be disrupted. The solute-solvent hydrogen bonding between glucose and water is shown in figure 12.

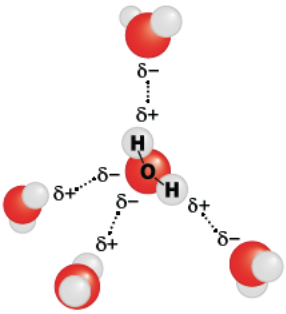


Figure 11 Hydrogen bonds between water molecules.

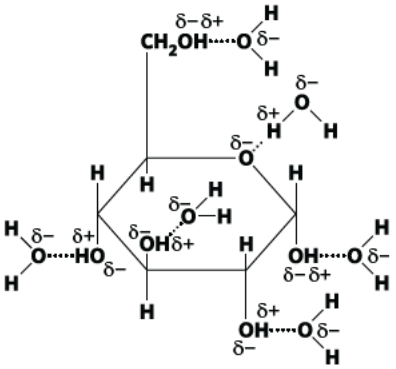


Figure 12 The five hydroxyl groups of glucose and the oxygen atom are able to form hydrogen bonds with many water molecules.

Therefore, in order for glucose to be dissolved, the hydrogen bonds between water molecules have been broken, the hydrogen bonds between glucose molecules have been broken and the hydrogen bonds between water and glucose molecules have formed.

This can be represented by the following chemical equation:

$$C_6H_{12}O_{6(s)} \xrightarrow{H_2O_{(l)}} C_6H_{12}O_{6(aq)}$$

It is important to note that the more polar the molecule is (usually due to more polar functional groups), the more soluble it is in water. Furthermore, while some molecules have some polar regions and some non-polar regions (e.g. carboxylic acids), as we learnt in lesson 6E, the influence of a polar functional group, or region, diminishes as the size of the non-polar region increases. This means that if we compare the molecular structures of vitamin C and vitamin D_2 as illustrated in figure 13, vitamin C will be water-soluble due to the presence of numerous polar groups. However, vitamin D_2 will be insoluble in water and soluble in non-polar solvents, even though it has a polar functional group.

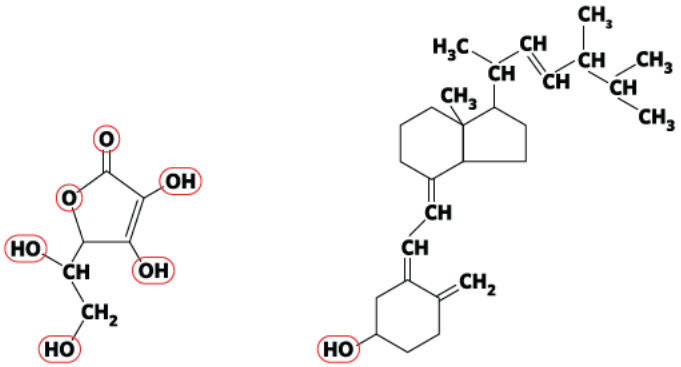


Figure 13 Comparison of vitamin C and vitamin D_2 with polar functional groups highlighted.

Ionisation of molecular compounds

Another way that molecular compounds are able to dissolve in water is through **ionisation**. Ionisation is the process by which ions are formed from a chemical reaction. The reason that some colecular compounds ionise is that some covalent bonds are so polar that when placed in a polar compound like water, they break into a positively and negatively charged ion (cation and anion respectively). One example is hydrogen chloride, HCl. HCl has a very electronegative chlorine atom bonded to a weakly electronegative hydrogen atom. When placed in water, the electronegative oxygen atom in water is attracted to the partially positive hydrogen atom in HCl, enough to tear apart the molecule as shown in figure 14. However, H^+ and Cl^-, do not exist isolated in solution but rather they become **hydrated**. This means the polar ions are dissolved in water with water molecules surrounding them as shown in figure 15. The force of attraction that keeps the ions in solution is called an **ion-dipole attraction** because the permanent dipole of water, either the partially positively charged hydrogen atoms or partially negatively charged oxygen atom, are attracted to the negative and positive ions. Therefore, overall we can say that the following chemical reaction has occurred.

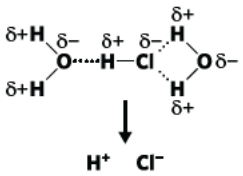


Figure 14 Ionisation of HCl in water.

$$HCl_{(g)} \xrightarrow{H_2O_{(l)}} H^+_{(aq)} + Cl^-_{(aq)}$$

These reactions will be explored in more detail in chapter 8. Some other compounds that are able to ionise like this are acids such as HNO_3, H_2SO_4 and CH_3COOH.

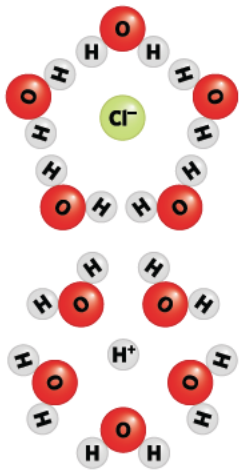


Figure 15 H^+ and Cl^- dissolved in solution.

Significance of the properties of water 2.1.2.3

OVERVIEW

The unique properties of water mean that it has very important applications in numerous domains including biology.

THEORY DETAILS

Water is often regarded as critical for life's success on Earth. As a result, discoveries of water on other planets are met with much excitement because there is a very close relationship between water and biological lifeforms.

Importance of water expanding when frozen

As we learnt previously in this lesson, as a result of the regular lattice arrangement that water molecules adopt at 0°C, water is able to expand. In a natural context, this is important because freezing conditions in water are generally inhospitable to life. Therefore, if it is very cold above the surface of the water, a layer of ice will form on the surface of the body of water as shown in figure 16. Ice, with its greater volume and lower density, floats and therefore forms a protective covering on the surface of the water to act as a thermal barrier. That is, it protects the life forms within the water below by maintaining the temperature of the water underneath it above 0°C.



Image: Standret/Shutterstock.com

Figure 16 Frozen lake.

In a household context, the expansion of water as it freezes has some negative effects. For examples, pipes that are located outside or in very cold areas can be damaged when water freezes. This is because the pipes, full of water, can be placed under pressure and eventually rupture as the water freezes and expands. This is why antifreeze is often added to any water-based compounds that are at risk of being frozen, such as in cars as shown in figure 17. That is, the water coolant in cars required for internal combustion engines to run effectively does not freeze, expand and cause the rigid space it is stored in to burst.



Image: Krasula/Shutterstock.com

Figure 17 Antifreeze is often added in cars in cold climates.

Importance of the solvent properties of water

In a biological context, water is a critical solvent for the chemical reactions needed to sustain life. This is why the body consists of approximately 60% water because aqueous solutions provide a perfect environment for chemical reactants to interact with each other, collide with each other and undergo chemical reactions. The circulatory system, illustrated in figure 18, uses the solvent properties of water to transport a number of different substances throughout the body, including but not limited to:

- Glucose (required for cells to produce energy)
- Waste matter (waste matter is produced from exercising muscles and is carried in the blood so it can be excreted)
- Ions (metal ions like Na^+, Cl^-, Mg^{2+} play a variety of important roles in the body)
- Hormones (hormones are released into the bloodstream to be carried around the body and have specific effects on the tissue/organ they are targeting)

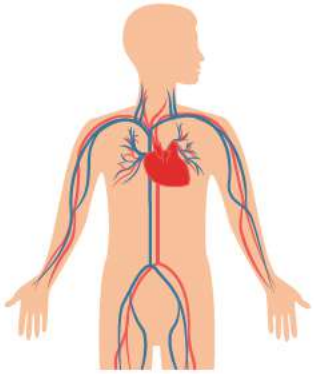


Figure 18 Circulatory system.

In the home, water is frequently used as a solvent. Many paints or varnishes contain solutes dissolved in water and soft drinks as well as other carbonated beverages contain CO_2 dissolved in water. Medications like paracetamol, can be taken dissolved in water as shown in figure 19.



Image: krolya25/Shutterstock.com

Figure 19 Medication dissolved in water.

In industry, water is frequently used as a solvent. For example, electroplating is a process where metal ions, like gold or silver, are coated onto a piece of metal. In this process, the gold and silver ions are aqueous and dissolved as shown in figure 20.

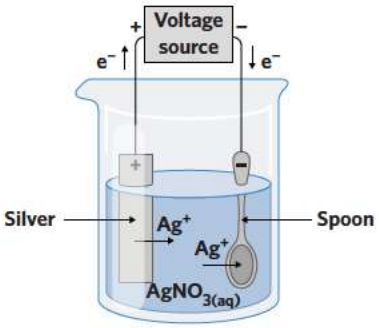


Figure 20 Industrial application of water as a solvent (electroplating).

Theory summary

- Water is unique with respect to its high melting point and boiling point when compared to other Group 16 hydrides.
- Water expands when frozen, becoming less dense.
- Water is able to dissolve other polar compounds, as well as molecular substances through hydrogen bonding and ionisation.
- The ability for water to act as a solvent is important in biological and industrial systems.

7A QUESTIONS

Theory review questions

Question 1

Which of the following is a Group 16 hydride?

A H_2O

B KH

Question 2

Hydrogen telluride is a Group 16 hydride. Its melting and boiling point is _____ than the melting and boiling point of water.

A greater

B less

Question 3

Which of the following represents water when it is in its solid state?

A



B



Question 4

What needs to be overcome before a substance is able to dissolve?

A Solute-solvent attractive forces

B Solute-solute attractive forces

Question 5

Which of the following molecules will dissolve in water by forming hydrogen bonds with water molecules?

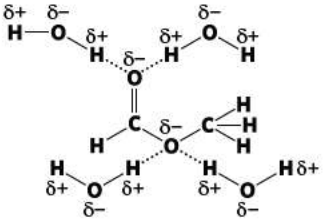
A HCl

B CH_3OH

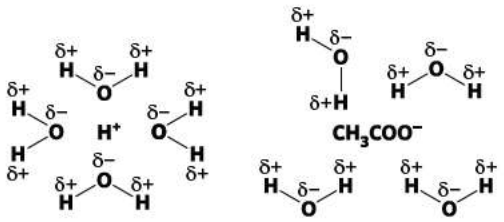
Question 6

The image that represents a compound that has dissolved in water through ionisation is:

A



B



Question 7

The most commonly used solvent in biological systems is

A polar.

B non-polar.

Question 8

Why is it easier for chemical species to react in solution compared to in their solid form?

A Chemical species are fixed in position in a solution and, therefore, are held in the perfect position for a reaction to occur.

B Chemical species are free to interact, collide and react with each other.

Question 9

If water is used as a solvent, which of the following compounds would not be able to dissolve to any significant extent?

A HNO_3

B $C_{12}H_{25}OH$

Question 10

Solutions consist of _____ particles which are evenly dispersed in a _____ mixture.

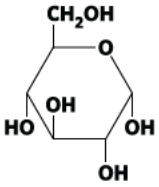
A solute, homogeneous

B solvent, inhomogeneous

Deconstructed questions

Use the following information to answer questions 11–13.

Glucose, $C_6H_{12}O_6$, is an important source of energy for the body and is transported through the bloodstream.



α - glucose

Question 11 (1 MARK)

Which of the following is a correct statement regarding the solubility of glucose?

A Glucose is soluble in non-polar solvents because 'opposites attract'.

B Glucose is soluble in polar solvents because 'like dissolves like'.

C Glucose is insoluble in polar solvents because glucose can form hydrogen bonds with the solvent.

D Glucose is insoluble in non-polar solvents because glucose can form hydrogen bonds with the solvent.

Question 12 (1 MARK)

In order for glucose to dissolve, which of the following correctly lists all of the forces that need to be broken and formed?

I Solute-solute forces broken

II Solute-solute forces formed

III Solvent-solvent forces broken

IV Solvent-solvent forces formed

V Solute-solvent forces broken

VI Solute-solvent forces formed

A I, IV, VI

B II, III, V

C I, III, VI

D I, III, V

Question 13 (3 MARKS)

Draw the interaction between glucose and water molecules. Explain why this enables solid glucose to dissolve in water and write out the chemical equation that represents this process.

Exam-style questions

Within lesson

Question 14 (1 MARK)

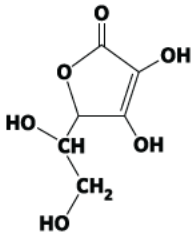
Which of the following substances will be able to dissolve in water without forming ions?

A Sulfuric acid (H_2SO_4)
B NItric acid (HNO_3)
C Hydrochloric acid (HCl)
D Methanol (CH_3OH)

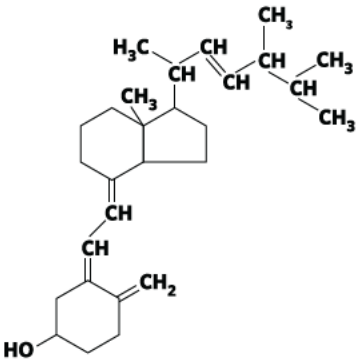
Question 15 (1 MARK)

Vitamins are important nutrients in the human body and can be either classified as either water-soluble or lipid-soluble. Lipids are non-polar compounds. Which of the following correctly lists whether a vitamin is water or lipid-soluble.

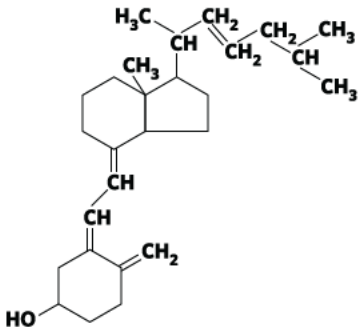
A **Name: Vitamin C**
Type: Lipid-soluble



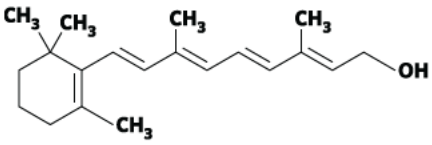
B **Name: Vitamin D_2**
Type: Water-soluble



C **Name: Vitamin D_3**
Type: Water-soluble



D **Name: Vitamin A**
Type: Lipid-soluble

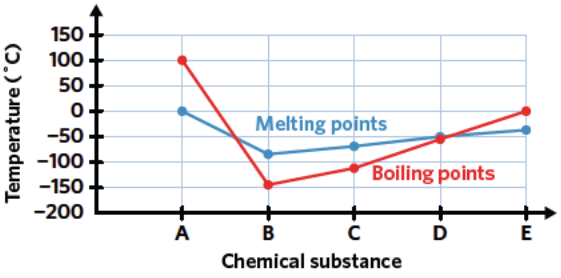


Question 16 (6 MARKS)

Water is one member of the Group 16 hydrides.

a What is a Group 16 hydride? (1 MARK)

Jessie is investigating the properties of Group 16 hydrides for a scientific poster. The following is a graph of the melting and boiling points of Group 16 hydrides that she has from her experiment, but unfortunately it is missing the labels.



b What is the name and the molecular formula of A. Explain why it has such a high melting and boiling point compared to other Group 16 hydrides. (3 MARKS)

c Excluding molecule A, why do the melting and boiling points increase from substance B to E? (2 MARKS)

Question 17 (10 MARKS)

The Titanic was the largest passenger cruise ship of its time in 1912.

a With reference to the volume and density of water in different states, why do icebergs float on water? Include a diagram in your answer. (3 MARKS)

b Why is it biologically important that ice floats on water? (2 MARKS)

When the Titanic was sinking, one officer called Chief Baker Joughin drank alcohol as the ship sank and was able to survive for a long time in the freezing water. It is rumoured that the alcohol in his blood acted as an antifreeze.

c Assume the alcohol is ethanol. Draw the hydrogen bonding interactions between ethanol and water molecules and explain why ethanol is soluble in water. (3 MARKS)

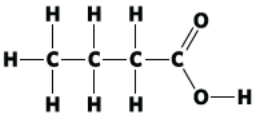
d Given water freezes due to regular hydrogen bonding arrangements between water molecules, suggest why ethanol could act as an antifreeze agent. (2 MARKS)

Multiple lessons

Question 18 (9 MARKS)

Carboxylic acids are a homologous series of compounds with the general formula $C_nH_{2n+1}COOH$.

a What is the name of the given carboxylic acid? (1 MARK)



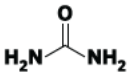
b Draw the structural formula for the largest product formed when pentan-1-ol reacts with propanoic acid in the presence of concentrated sulfuric acid. What type of functional group is formed and what is the name of this compound? (3 MARKS)

c Will this compound be soluble in water? With reference to a diagram, explain why or why not. (3 MARKS)

d Ethanoic acid is another example of a carboxylic acid. It can dissolve in water through ionisation. Write out the chemical equation that represents this process and draw the ions as they would exist in solution. (2 MARKS)

Question 19 (9 MARKS)

In mammals, one source of energy that is sometimes used is the oxidation of amino acids. During this process, ammonia (NH_3) is produced which is highly toxic and so it is immediately converted into urea as shown:



a Is urea soluble in water? Justify your answer with reference to a diagram. (3 MARKS)

b Based on its solubility properties, explain how urea can be transported in the blood stream and then excreted in urine. (2 MARKS)

Urea plays an important role in water reabsorption in the kidney by being more heavily concentrated in the kidney than the urine. This is because water tends to go from regions of low solute concentration to regions of high solute concentration (a process called osmosis).

c In this scenario, is urea the solute or solvent? (2 MARKS)

d Urea cycle diseases lead to a buildup of ammonia (NH_3), a toxic compound, in the body. The main symptoms are neurological and include seizures, coma and, in severe cases, death. If ammonia is produced in the liver, how is it able to affect the brain? (2 MARKS)

Key science skills

Question 20 (5 MARKS)

Margaret is testing whether certain compounds can be dissolved in water. She believes that she can do this visually by placing different solid compounds into water, and then observing whether they are visible or not after 30 seconds.

a What type of data is being collected? (1 MARK)

b What is the independent variable in this experiment? (1 MARK)

For one experiment, Margaret times how long it takes for 5 g of a solid substance to dissolve. She places 5 g of substance A into a beaker and 5 g of substance B into another beaker, each with the same volume of water. However, she stirs substance A but not substance B. Both substances are completely dissolved after exactly 20 seconds.

c Is this experiment valid? Explain why or why not. (2 MARKS)

d Suggest one improvement to the experimental design to improve its validity. (1 MARK)

Exam-style question hints

14: Molecular substances are able to dissolve in water either through hydrogen bonding or ionisation. **15:** Molecules with a polar functional group are not always water-soluble. **16a:** Group 16 elements include oxygen, sulphur and tellurium. **16b:** Water is the only Group 16 hydride able to form hydrogen bonds. **16c:** Dispersion forces increase as molecules grow larger. **17a:** Water is able to form a regular, crystal-like hydrogen bonding arrangement when frozen. **17b:** Ice can act as a barrier between the cold temperatures above the water and the water below. **17c:** Hydrogen bonding interactions occur between a hydrogen atom bonded to an electronegative atom and another electronegative atom. **17d:** Antifreeze agents disrupt the stable hydrogen bonding arrangement of water to prevent it from freezing. **18a:** The name of a carboxylic acid depends on the number of carbon atoms in the carbon chain. **18b:** Esters are produced from the reaction between a carboxylic acid and an alcohol. **18c:** Not all compounds with polar functional groups are soluble in water. **18d:** CH_3COOH ionises in a similar manner to HCl. **19a:** Hydrogen bonds can occur when hydrogen is bonded to either fluorine, oxygen or nitrogen. **19b:** Blood and urine consist mostly of water. **19c:** Osmosis refers to the movement of the solvent which in this case is water. **19d:** Ammonia is very soluble in water. **20a:** Data can either be numbers or observations. **20b:** The independent variable is the variable that is changed by the experimenter. **20c:** Validity refers to whether an experiment accurately measures the aim **20d:** All variables except the independent variable should be controlled in an experiment.

7B THERMAL PROPERTIES OF WATER

The ability of water to absorb large amounts of thermal energy is important for the survival of biological organisms.

7A Properties of water | 7B Thermal properties of water | 7C Dissolution | 7D Concentration | 7E Solubility | 7F Precipitation reactions

Study design dot point

- specific heat capacity and latent heat including units and symbols, with reference to hydrogen bonding to account for the relatively high specific heat capacity of liquid water, and significance for organisms and water supplies of the relatively high latent heat of vaporisation of water

Key knowledge units

Specific heat capacity	2.1.2.1
Latent heat	2.1.2.2

Key terms and definitions

Heat capacity relationship between the heat absorbed by a substance and its temperature change

Specific heat capacity energy (J) required to raise the temperature of 1 gram of a given substance by 1°C

Latent heat of fusion amount of energy required to convert 1 mol of a substance from its solid state to its liquid state at the melting point of the substance

Latent heat of vaporisation amount of energy required to convert 1 mol of a substance from its liquid state to its gaseous state at the boiling point of the substance

Lesson links

This lesson builds on:

- 7A Properties of water

The thermal properties of water are based on water's molecular structure.

Specific heat capacity 2.1.2.1

OVERVIEW

Heat capacity is a measure of how well a substance is able to absorb and store thermal energy.

THEORY DETAILS

Heat capacity is a general term used to describe the amount of thermal energy a substance needs to absorb for a temperature change to occur. More specifically, it is the amount of thermal energy required to raise the temperature of a substance by a certain number of degrees. For example, if Substance A and Substance B have different heat capacities, when the same amount of thermal energy is applied to each substance, they will experience different temperature changes. Applying this concept to a real world scenario, if a piece of metal and a glass of water were left in the sun, the metal would feel the hottest to touch after a set amount of time, even though both received the same amount of thermal energy as shown in figure 1. This is due to the higher heat capacity of water compared to the lower heat capacity of the piece of metal.

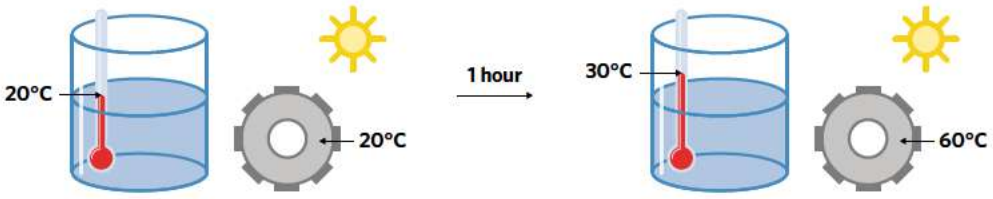
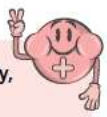


Figure 1 Demonstration of heat capacity.

While heat capacity can be a useful term to describe the differences in thermal properties between compounds, it is very general and has no set units. Therefore, scientists have created another term called **specific heat capacity** which describes the amount of energy required (measured in joules, J) to raise the temperature of 1 gram of substance by 1°C. Therefore, specific heat capacity, which is given the symbol c, has the units $J\,g^{-1}{}^{o}C^{-1}$.



Specific heat capacity, c, is given the units $J\,g^{-1}{}^{o}C^{-1}$ and since a temperature change of 1°C is the same as a temperature change of 1K, the units can also be expressed as $J\,g^{-1}\,K^{-1}$.

Each substance has a different heat capacity and table 1 displays some of the heat capacities of common substances.

Table 1 Specific heat capacities of different substances.

Substance	Specific heat capacity (J $g^{-1}{}^{o}C^{-1}$)
Water	4.18
Lithium	3.56
Ethanol	2.46
Sodium	1.23
Aluminium	0.90
Concrete	0.88

As can be seen, there are a range of specific heat capacities across metals, molecular compounds and composite materials. It is important to note that water has an unusually high specific heat capacity of 4.18 J $g^{-1}{}^{o}C^{-1}$ which is consistent with the excpetional properties of water explored in lesson 7A. This means that it takes 4.18 joules of energy to raise the temperature of 1 gram of water by 1°C.



The specific heat capacity of water is an important constant and the value can be found in the VCE chemistry data book.

The high specific heat capacity of water is due to water's hydrogen bonding properties and is a reflection of how much thermal energy the intermolecular bonds are able to absorb before breaking. In the case of water, the strong hydrogen bonds between water molecules enable a large amount of thermal energy to be stored, as opposed to molecules or compounds with weaker intermolecular forces. In a biological context, the high specific heat capacity of water means that it can maintain a relatively constant temperature because it requires a large amount of thermal energy before it increases in temperature. This allows biological organisms to maintain constant internal temperatures which is vital for survival.

Calculations involving specific heat capacity

The specific heat capacity of water is an important constant in chemistry because it defines the amount of thermal energy required to raise the temperature of water by a certain number of degrees. Since water is a liquid, it is often measured by volume in mL, or L. This means that the volume of water must be converted to a mass in grams to perform calculations. In the data book, the density of water is given as 0.997 g mL^{-1} at 25°C which means that at 25°C 1 mL of water weighs 0.997 g. Therefore, to convert from volume in mL to mass in grams, the volume of water is multiplied by 0.997. For example, 2.50 L of water, which is 2500 mL, weighs 2500 mL × 0.997 g mL^{-1} = 2492.5 g = 2.49 kg.



The VCE chemistry data book specifies that the density of water is 0.997 g mL^{-1} at 25°C, not 1.000 g mL^{-1}.

For the most part, the approximation 1 mL = 1 g is sufficient but for larger volumes of water, there will be inaccuracies and therefore, it is best to use 1 mL = 0.997 g.

After calculating the mass of water present, the heat energy required to raise the temperature of the water by a certain number of degrees can be calculated using the following equation:

$$q = m \times c \times \Delta T$$

q is the energy that is transferred (joules), m is the mass of water (grams), c is the specific heat capacity of water (4.18 J $g^{-1}{}^{o}C^{-1}$) and ΔT is the change in temperature of the water. As we are only measuring the change in temperature of the water, we can use either celsius or kelvin as our unit of temperature. The change in temperature is calculated using the following equation:

$$\Delta T = T_{final} - T_{initial}$$



$q = m \times c \times \Delta T$ is found in the VCE chemistry data book and can be remembered with $q = mc\Delta t$.



Worked example 1

Elizabeth is investigating how much heat energy is required to make a cup of tea. She knows that each cup of tea requires 250 mL of water and the initial temperature of the water is 25°C. She also knows that the final temperature of water required is 100°C and that all heat energy in the kettle is transferred to the water. How much energy is required for Elizabeth to make a cup of tea?

What information is presented in the question?

Volume of water: 250 mL

Initial temperature: 25°C

Final temperature: 100°C

What is the question asking us to do?

Calculate the heat energy required to raise the temperature of 250 mL of water from 25°C to 100°C.

What strategy(ies) do we need in order to answer the question?

1. Determine the mass of water present.
2. Calculate the temperature change of the water.
3. Using the specific heat capacity of water, substitute values into $q = m \times c \times \Delta T$ to calculate the amount of heat energy required.

Answer

The mass of water present is $250\ \text{mL} \times 0.997\ \text{g mL}^{-1} = 249\ \text{g}$

The temperature change of water is calculated by:

$\Delta T = T_{final} - T_{initial}$

$= 100°\text{C} - 25°\text{C} = 75°\text{C}$

The specific heat capacity of water is $4.18\ \text{J g}^{-1}\text{°C}^{-1}$

Substitute these values into the equation:

$q = m \times c \times \Delta T$

$q = 249\ \text{g} \times 4.18\ \text{J g}^{-1}\text{°C}^{-1} \times 75°\text{C} = 78\,061.5\ \text{J}$

$= 78\ \text{kJ}$

Hence, 78 kJ of energy is required to make a cup of tea.

The specific heat capacity of water can also be used to calculate the temperature change that a certain volume/mass of water will experience when a certain amount of heat energy is applied.



Worked example 2

When combusted, ethanol releases 29.0 kJ per gram. Calculate the final temperature of 350 mL of water heated by the combustion of 1.0 g of ethanol. Assume the temperature of the water is initially at 25°C and all heat energy released from the combustion of ethanol is transferred to the water.

What information is presented in the question?

Volume of water: 350 mL

Initial temperature: 25°C

Heat energy released: 29 kJ

What is the question asking us to do?

Calculate the final temperature of 350mL of water initially at 25°C when 1.0 g of ethanol is combusted.

What strategy(ies) do we need in order to answer the question?

1. Determine the mass of water present.
2. Convert the amount of heat energy expressed in kJ to J.
3. Using the specific heat capacity of water, substitute the values into $q = m \times c \times \Delta T$ to calculate the temperature change of the water.
4. Determine the final temperature of the water by using the equation $\Delta T = T_{final} - T_{initial}$.

Answer

The mass of water present is $350\ \text{mL} \times 0.997\ \text{g mL}^{-1} = 349\ \text{g}$.

1.0 g of ethanol releases 29.6 kJ = 29.6 × 1000 = 29 600 J.

The specific heat capacity of water is $4.18\ \text{J g}^{-1}\text{°C}^{-1}$.

Substitute these values into the equation:

$q = m \times c \times \Delta T$

$29\,600\ \text{J} = 349\ \text{g} \times 4.18\ \text{J g}^{-1}\text{°C}^{-1} \times \Delta T$

$$\Delta T = \frac{29\,600\ \text{J}}{349\ \text{g} \times 4.18\ \text{g}^{-1}\text{°C}} = 20°\text{C}$$

$\Delta T = T_{final} - T_{initial}$

$20°\text{C} = T_{final} - 25°\text{C}$

$T_{final} = 45°\text{C}$

Therefore, the final temperature of the water is 45°C.



Latent heat 2.1.2.2

OVERVIEW

Water has a relatively high latent heat which enables it to absorb large amounts of thermal energy before changing state.

THEORY DETAILS

As a solid is heated under constant pressure, its temperature will increase until it has reached its melting point. At this point, the solid doesn't immediately change state into a liquid. Energy is continually absorbed by the solid as it melts without raising the temperature. Once enough energy has been absorbed, the state change from solid to liquid can occur. This concept is known as the **latent heat of fusion**. That is, energy known as the latent heat of fusion is required for the following state change:

$A_{(s)} \rightarrow A_{(l)}$

Similarly, as a liquid is heated under constant pressure, its temperature will increase until it has reached its boiling point. At this point, the liquid doesn't immediately change state into a gas. Energy is continually absorbed by the liquid as it boils without raising the temperature. Once enough energy has been absorbed, the state change from liquid to gas can occur. This concept is known as the **latent heat of vaporisation**. That is, energy known as the latent heat of vaporisation is required for the following state change:

$A_{(l)} \rightarrow A_{(g)}$

Using water as an example, the latent heats of fusion and vaporisation can be visualised in the graph in figure 2 as the regions where energy continues to be absorbed as the state change occurs, but there is no temperature change.

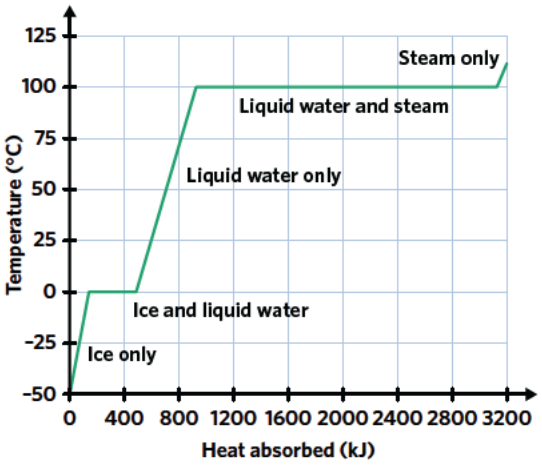


Figure 2 Graph of the latent heat of fusion and vaporisation of water.

The latent heats of fusion and vaporisation also reflect the strength of intermolecular forces, which explains why water has high latent heats compared to other molecules, as shown in table 2.

Table 2 Comparison of the latent heats of different substances.

Substance	Latent heat of fusion (kJ/mol)	Latent heat of vaporisation (kJ/mol)
Water (H_2O)	6.0	44.0
Hydrogen (H_2)	0.12	0.90
Oxygen (O_2)	0.44	6.82

The latent heat of fusion for water is 6.0 kJ mol^{-1}. This means that 6.0 kJ of energy is required to convert 1 mol of ice to liquid water. This energy is required to disrupt the hydrogen bonds between water molecules that are held in a regular lattice at freezing point.

The latent heat of vaporisation for water is 44.0 kJ mol^{-1}. This means that it requires 44.0 kJ of energy to convert 1 mol of water from a liquid to a gas. This energy is required to completely disrupt the hydrogen bonding arrangement between water molecules so these molecules are free to move about and fill any available space as a gas.

Just like calculations involving the specific heat capacity, we can also calculate the amount of energy required for a state change to occur. In this case, the formula is:

$q = n \times L$

q is the heat energy (kJ), n is the amount of substance present and L is the latent heat of fusion or vaporisation (kJ mol^{-1}).

Worked example 3

Calculate the heat energy required to convert 845 g of ice to liquid water at water's melting point.

What information is presented in the question?

Water is melting.

Mass of water: 845 g

What is the question asking us to do?

Use the latent heat of fusion to calculate how much heat energy is required to melt 845 g of ice.

What strategy(ies) do we need in order to answer the question?

1. Calculate the amount (in mol) of water present.
2. Using the value for the latent heat of fusion for water, calculate the heat energy with the equation $q = n \times L$.

Answer

$$n(H_2O) = \frac{m}{M} = \frac{845}{(2 \times 1.0) + 16.0} = 46.9 \text{ mol}$$

$$L(H_2O) = 6.0 \text{ kJ mol}^{-1}$$

$$q = n \times L = 46.9 \text{ mol} \times 6.0 \text{ kJ mol}^{-1} = 2.8 \times 10^2 \text{ kJ}$$

Therefore, 2.8×10^2 kJ of heat energy is required to melt 845 g of water.

Worked example 4

Calculate the heat energy required to convert 450 g of liquid water to steam at water's boiling point.

What information is presented in the question?

Water is boiling.

Mass of water: 450 g

What is the question asking us to do?

Use the latent heat of vaporisation to calculate how much heat energy is required to boil 450 g of water.

What strategy(ies) do we need in order to answer the question?

1. Calculate the amount (in mol) of water present.
2. Using the value for the latent heat of vaporisation for water, calculate the heat energy with the equation $q = n \times L$.

Answer

$$n(H_2O) = \frac{m}{M} = \frac{450}{(2 \times 1.0) + 16.0} = 25.0 \text{ mol}$$

$$L(H_2O) = 44.0 \text{ kJ mol}^{-1}$$

$$q = n \times L = 25.0 \text{ mol} \times 44.0 \text{ kJ mol}^{-1} = 1.10 \times 10^3 \text{ kJ}$$

Therefore, 1.10×10^3 kJ of heat energy is required to boil 450 g of water.

Significance of water's high latent heat of vaporisation

In a biological context, organisms use the high latent heat of vaporisation for thermoregulation which is the maintenance of a relatively stable body temperature. Sweating is important to avoid overheating because the water on the surface of the skin can absorb a large amount of heat energy from the body before evaporating, as shown in figure 3. If the latent heat of vaporisation of water was a smaller quantity, the process of sweating would be less effective. Therefore, the high latent heat of vaporisation means that large bodies of water evaporate gradually as water absorbs and retains a large amount of heat energy.

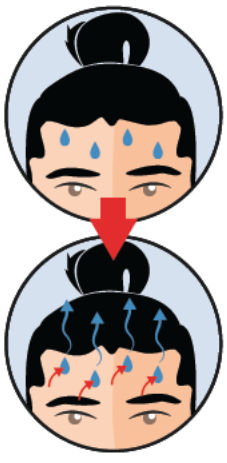


Figure 3 The process of sweating.

Theory summary

- The heat capacity of a substance is the amount of heat energy required to raise the temperature of a substance by a certain number of degrees.
- The specific heat capacity of water is 4.18 J $g^{-1}{}^{o}C^{-1}$ and is used in the equation $q = m \times c \times \Delta T$.
- The latent heat of fusion is the amount of heat energy required to convert a substance from its solid state to its liquid state.

- The latent heat of vaporisation is the amount of heat energy required to convert a substance from its liquid state to its gaseous state.
- Calculations involving latent heat use the equation $q = n \times L$.
- Water has a high latent heat of fusion (6.0 kJ mol^{-1}) and a high latent heat of vaporisation (44.0 kJ mol^{-1}).

7B QUESTIONS

Theory review questions

Question 1

When two different substances absorb the same amount of thermal energy, they will experience different temperature changes.

A True

B False

Question 2

Specific heat capacity is measured in

A kJ mol^{-1}.

B J $g^{-1}{}^{\circ}C^{-1}$.

Question 3

Latent heat is measured in

A kJ mol^{-1}.

B J $g^{-1}{}^{\circ}C^{-1}$.

Question 4

Which of the following equations represents water's latent heat of vaporisation?

A $H_2O_{(s)} \rightarrow H_2O_{(l)}$

B $H_2O_{(l)} \rightarrow H_2O_{(g)}$

Question 5

Which of the following equations represents water's latent heat of fusion?

A $H_2O_{(s)} \rightarrow H_2O_{(l)}$

B $H_2O_{(l)} \rightarrow H_2O_{(g)}$

Question 6

In order to perform calculations involving water's latent heat, which of the following equations should be used?

A $q = m \times c \times \Delta T$

B $q = n \times L$

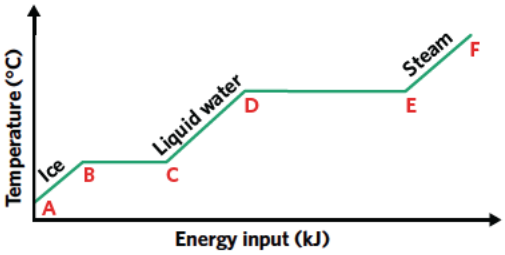
Question 7

In order to perform calculations involving water's specific heat capacity, which of the following equations should be used?

A $q = m \times c \times \Delta T$

B $q = n \times L$

Use the following information to answer questions 8–10.



Question 8

How much energy is required to go from point B to C?

A 6.0 kJ mol^{-1}

B 44.0 kJ mol^{-1}

Question 9

How much energy is required to go from point D to E?

A 6.0 kJ mol^{-1}

B 44.0 kJ mol^{-1}

Question 10

Between point C and D, how much energy is required to raise the temperature of 1 g of water by 1°C?

A 44.0 kJ

B 4.18 kJ

Deconstructed questions

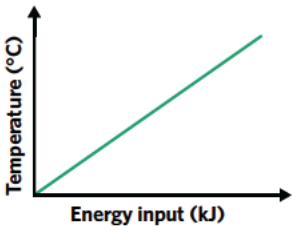
Use the following information to answer questions 11–13.

Suzie wants to determine how much energy is required to convert 150 g of ice at 0°C to steam at 100°C.

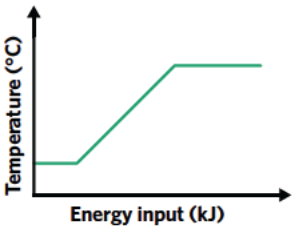
Question 11 (1 MARK)

Which of the following graphs best represents the process of converting ice to steam?

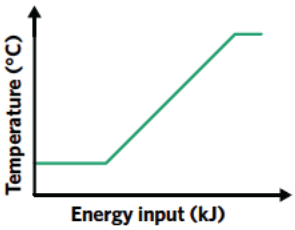
A



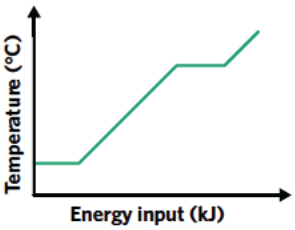
B



C



D



Question 12 (1 MARK)

In order to determine how much heat energy is required, Suzie needs to use:

I Latent heat of fusion

II Latent heat of vaporisation

III Specific heat capacity of water

A III only

B I, II

C II, III

D I, II, III

Question 13 (4 MARKS)

Calculate how much heat energy Suzie requires. Show all working.

Exam-style questions

Within lesson

Question 14 (1 MARK)

What is the energy required to convert 250 mL of water at 25.0°C to steam at 100°C?

A 78.1 kJ
B 78.4 kJ
C 687 kJ
D 720 kJ

Question 15 (1 MARK)

What is the final temperature of 50 g of water at 27°C after it is heated with 12 kJ of energy?

A 28°C
B 84°C
C 62°C
D 93°C

Question 16 (6 MARKS)

The high latent heat and specific heat capacity of water has many practical applications.

a Define specific heat capacity and latent heat. (2 MARKS)

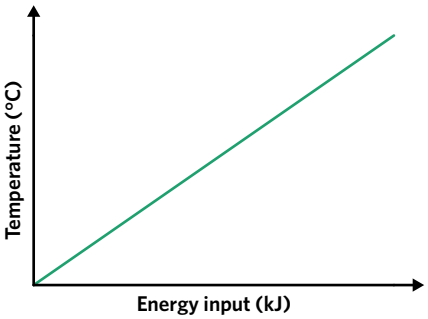
b With reference to the intermolecular forces present, explain why water has high values for its specific heat capacity and latent heat. (2 MARKS)

Water is used as a coolant in electronic products to remove excess heat.

c What property of water makes it desirable to remove excess heat in electronics? Justify your response. (2 MARKS)

Question 17 (10 MARKS)

Walter is investigating the thermal properties of water for a scientific project. He develops a graph that depicts the relationship between the amount of energy absorbed by a sample of water and the resulting temperature increase of the water. He predicts that graph will look like this:



The real relationship, however, is not linear.

a Why would Walter predict that there is a linear relationship between the temperature of water and energy input? (2 MARKS)

b In reality, there are two regions where the temperature does not change but energy is absorbed by the water. Identify what these regions are called and their values for water. (2 MARKS)

c These regions also differ in size. Which one is larger? Explain. (2 MARKS)

d What is the energy required to convert 25 g of ice to liquid water at water's melting point? (2 MARKS)

e What is the energy required to convert 25 g of liquid water to steam at water's boiling point? (2 MARKS)

Multiple lessons

Question 18 (9 MARKS)

The concept of specific heat capacity can also be applied to molecules other than water.

a Draw the structural formula of ethanol. What is the functional group present? (2 MARKS)

b Explain why ethanol has a relatively high specific heat capacity that is smaller than water's. (3 MARKS)

The combustion of 1.0 g of ethanol releases 29.6 kJ. 300 mL of water at 20°C is heated by the combustion of 3.0 g of ethanol.

c What is the final temperature of the water? Assume all the energy released from the combustion of ethanol is transferred to the water. (4 MARKS)

Question 19 (6 MARKS)

When metals are left in the sun, they become very hot quickly. This means that on a hot day, the exterior of a car can sometimes fry an egg.

a Describe metallic bonding. (2 MARKS)

b Explain why metals are good conductors of heat. (2 MARKS)

c Explain why the property of heat conduction leads to a low specific heat capacity. (2 MARKS)

Key science skills

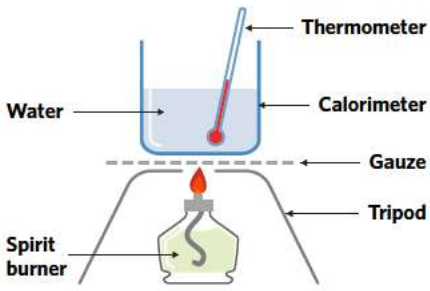
Question 20 (4 MARKS)

Calorimetry is the process of measuring the energy released by the combustion of a substance by determining the temperature change of the water. An example of a calorimeter is shown.

a What type of data is collected in a calorimetry experiment? (1 MARK)

b Provide an exmple of a systematic error in this experimental setup. (1 MARK)

c Suggest two ways to improve the experiment. (2 MARKS)



Exam-style question hints

14: Water's latent heat of vaporisation is 44.0 kJ mol^{-1}. **15:** $q = m \times c \times \Delta T$ calculates the temperature change, not the final temperature. **16a:** The units of latent heat and specific heat capacity reflect what is measured. **16b:** Intermolecular bonding affects the amount of heat required. **16c:** Excess heat can be transferred to another substance. **17a:** Formulas describe relationships between variables. **17b:** As a result of latent heat, energy is absorbed but there is no temperature change. **17c:** State changes require different amounts of energy. **17d:** The latent heat of fusion of water is 6.0 kJ mol^{-1}. **17e:** The latent heat of vaporisation of water is 44.0 kJ mol^{-1}. **18a:** Ethanol is an alcohol. **18b:** Intermolecular bonds affect specific heat capacity. **18c:** $q = m \times c \times \Delta T$ calculates the temperature change, not the final temperature. **19a:** Metals have a unique form of intermolecular bonding. **19b:** Metals have a sea of delocalised electrons. **19c:** Heat conduction and specific heat capacity are related. **20a:** Data can either relate to observations or numbers. **20b:** Systematic errors affect the experiment to the same extent each time. **20c:** Think about ways to reduce the main systematic error.

7C DISSOLUTION

In this lesson, we will learn about how the properties of water allow other substances to be dissolved in water.

7A Properties of water | 7B Thermal properties of water | 7C Dissolution | 7D Concentration | 7E Solubility | 7F Precipitation reactions

Study design dot point

- the comparison of solution processes in water for molecular substances and ionic compounds

Key knowledge units

Dissolution of ionic substances in water	2.1.3.2
Solubility tables	2.2.4.1
Dissolution of molecular substances in water	2.1.3.1

Key terms and definitions

Dissociate (ions) process where ions break away from their lattice

Dissociation process by which a molecule separates into smaller particles

Dissolve substance becomes incorporated into a liquid to form a solution

Dissolution substance dissolves in another substance

Soluble able to be dissolved in a given solvent

Insoluble unable to be dissolved in a given solvent

Lesson links

This lesson builds on:

- 4A Ionic compounds
  The charged nature of ionic compounds allows them to interact with polar molecules.
- 7A Properties of water
  The polar nature of water allows different substances to dissolve in it.
- 5B Properties of covalent compounds
  The polarity of a molecular compound will affect its ability to dissolve in water.

Dissolution of ionic substances in water 2.1.3.2

OVERVIEW

The charged nature of ionic compounds allows them to dissociate in water.

THEORY DETAILS

As we saw in lesson 4A, ionic compounds are made up of positively and negatively charged ions known as cations and anions. As a result, ionic compounds can interact with substances, like water, that consist of polar molecules. This causes the ionic compounds to **dissociate** into their individual ions.

The most familiar example of this is what occurs when sodium chloride (NaCl), commonly known as table salt, mixes with water. When a sample of $NaCl_{(s)}$ is added to water, the polar water molecules start to orientate themselves in a way that allows either the partially positive or partially negative side of the water molecule to interact with the oppositely charged ion in the ionic compound that is exposed to water. The ion-dipole interactions between the water molecule and ions can be strong enough to pull the ions from the lattice structure, and to break the intermolecular interactions between water molecules. These ions become surrounded by water molecules and are described as being hydrated. This whole process, also referred to as **dissociation**, is shown in figure 1.

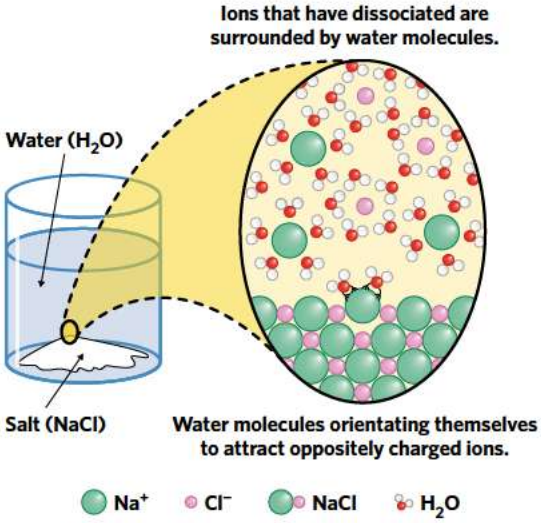


Figure 1 Dissociation of sodium chloride.

We can appreciate that no new products are formed; just that the ions in the ionic compound are dissociated from the lattice and can therefore move freely in the solution. In our example with NaCl, we can express this as

$$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$

As described by the equation, the solid sodium chloride has now dissociated in water to form sodium and chloride ions in a solution. We don't need to show water as part of this process; we can imply this from the change in phase of sodium chloride from solid to aqueous. As such, we can say that sodium chloride has **dissolved** in water. The process of dissolving is also referred to as **dissolution**. Ionic compounds that can dissolve in water are referred to as being water **soluble.**

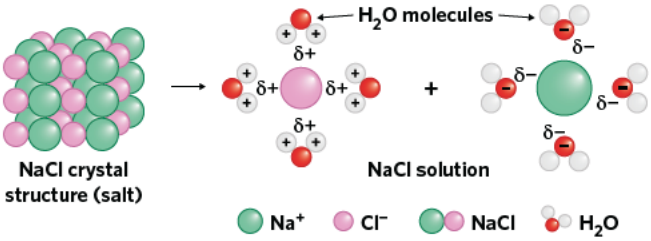


Figure 2 Interaction between water and ionic compounds.

Worked example 1

Solid potassium chloride was placed in water. Write an equation to show what would occur.

What information is presented in the question?

The name of the ionic compound, potassium chloride.

What is the question asking us to do?

Write an equation to represent what happens when potassium chloride is added to water.

What strategy(ies) do we need in order to answer the question?

1. Write the formula for potassium chloride.
2. Write the whole chemical equation to show potassium chloride dissociation into its ions.

Answer

The ionic formula for potassium chloride is KCl.

$KCl_{(s)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)}$

Although it may seem similar, dissolution and melting are two very different processes. Melting only requires one substance and results in a change in phase of the same substance. In contrast, dissolution involves two substances and results in a solution that is a mixture of both.

Not all ionic substances are able to dissolve in water and are described as being water **insoluble**. This is because the energy required to break the ions away from the lattice is more than what is released by the dissociation process.

Solubility tables 2.2.4.1

OVERVIEW

The solubility table helps us predict the solubility of ionic compounds.

THEORY DETAILS

A solubility table can be used to determine whether a substance is able to dissolve in water. The table shows a list of ions that easily dissolve in water at 25°C, as well as any exceptions. This allows us to predict interactions between an ionic compound and water.

A basic solubility table is shown in table 1, which has three columns showing the following:

1 The ion in question.
2 Whether this ion is generally soluble in water or not.
3 The exceptions to the rule.

As an example of how to read the solubility table, chloride ions (row 4) are shown to be soluble in water. However, as shown in the exceptions column, chloride ions that form a compound with Ag^+ ions to form AgCl are not soluble in water. These rules apply for all of the ions examined in the solubility table.

Table 1 Solubility table of common ions.

Ions	Solubility	Exceptions
Sodium ions – Na^+	Soluble	None
Potassium ions – K^+	Soluble	None
Ammonium salts – NH_4^+	Soluble	None
Chloride ions – Cl^-	Soluble	Insoluble in a compound with Ag^+, Hg^{2+}, Pb^{2+}
Iodide ions – I^-	Soluble	Insoluble in a compound with Ag^+, Hg^{2+}, Pb^{2+}
Nitrates – NO_3^-	Soluble	None
Perchlorate – ClO_4^-	Soluble	None
Ethanoate – CH_3COO^-	Soluble	None
Sulfate – SO_4^{2-}	Soluble	Insoluble in a compound with Ca^{2+}, Ba^{2+}, Sr^{2+}, Ag^+, Hg^{2+}, Pb^{2+}
Carbonate – CO_3^{2-}	Insoluble	Soluble in a compound with group IA* and NH_4^+
Phosphate – PO_4^{3-}	Insoluble	Soluble in a compound with group IA* and NH_4^+
Hydroxides – OH^-	Insoluble	Soluble in a compound with group IA*, Ca^{2+}, Ba^{2+}, Sr^{2+}
Sulfide – S^{2-}	Insoluble	Soluble in a compound with group IA*, IIA** and NH_4^+

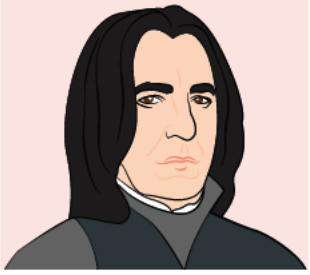
* Group IA metals include Li, Na, K, Rb, Cs, and Fr ** Group IIA metals include Be, Mg, Ca, Sr, Ba and Ra

The molecules included in the SNAPE rule are all highly soluble in water, and this therefore proves to be a good reference point for determining whether a compound is soluble.



A useful way of remembering whether an ion is likely to be soluble in water or not is through the use of the SNAPE rule. This rule states that if a compound contains one of the following ions in it, then it will be soluble in water:

- **S**odium (Na^+)
- **N**itrate (NO_3^-)
- **A**mmonium (NH_4^+)
- **P**otassium (K^+)
- **E**thanoate (CH_3COO^-)



Worked example 2

Using a solubility table, determine which of the following compounds will be soluble in water: AgI, Na_2S, KCl.

What information is presented in the question?

The identity of the molecules being analysed.

What is the question asking us to do?

Determine which of the molecules will be soluble in water.

What strategy(ies) do we need in order to answer the question?

1. Find the ions in the compounds in the solubility table.
2. Determine which of the ions are generally soluble in water.
3. Look for any exceptions to solubility.

Answer

AgI – Iodide ions (I^-) are generally soluble in water, however there is an exception when in a compound with silver ions (Ag^+), and therefore this compound will be water insoluble.

Na_2S – Although sulfide ions (S^{2-}) are generally insoluble, compounds containing sodium (Na^+) are always water soluble. Therefore Na_2S is considered soluble.

KCl – Both potassium and chloride are generally water soluble with no exception.

Therefore, both Na_2S and KCl are soluble in water.

The concept of insoluble ionic compounds will be further investigated in a later lesson.

Dissolution of molecular substances in water 2.1.3.1

OVERVIEW

The ability of molecular substances to dissolve depends on their polarity.

THEORY DETAILS

Due to water's polar nature, only polar substances have the capacity to dissolve in water. As we saw in chapter 5, there are many different molecular compounds that are polar. For example, alcohols, carboxylic acids and even esters can be considered as being polar due to the polar covalent bonds between either O-H or C=O present in the compound. These functional groups can form either hydrogen bonds or dipole dipole interactions with water, thus allowing them to dissolve in water as shown in figure 3.

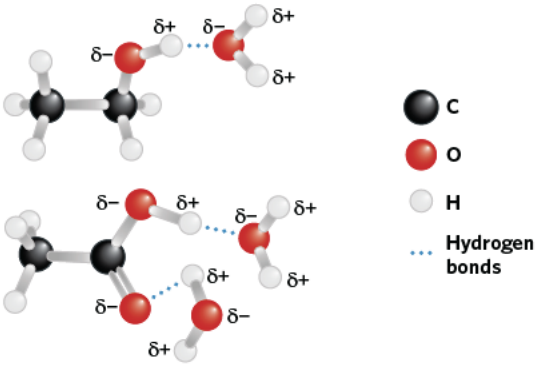


Figure 3 Interaction between water and polar molecular compounds.

As we saw in lesson 7A, in order for these molecular substances to dissolve, the hydrogen bonding between water molecules and the intermolecular bonds holding together the molecular substances need to break, allowing for the interaction between the molecules and water. To represent the dissolution of molecular substances in water, we only need to show the change in state. In the example of the dissolution of ethanol, we can represent this as:

$$C_2H_5OH_{(l)} \xrightarrow{H_2O_{(l)}} C_2H_5OH_{(aq)}$$



Worked example 3

A student has access to propane and propan-1-ol. Which of the two substances would be able to dissolve in water? Explain.

What information is presented in the question?

The names of two substances, propane and propan-1-ol.

What is the question asking us to do?

Determine which of the two will be soluble and explain why.

What strategy(ies) do we need in order to answer the question?

1. Determine the polarity of both compounds.
2. Identify the compound(s) that can dissolve in water.

Answer

Propane is a non-polar substance whereas propan-1-ol contains a polar hydroxyl group. As a result, propanol will be able to dissolve by forming hydrogen bonds with water whereas propane will not be able to dissolve in water.

Not all alcohols, carboxylic acids and esters are soluble in water. As we know, an increase in the non-polar component of molecular compounds reduces the impact of polar functional groups in the compound. Therefore, as the non-polar component of a molecular substance increases, the substance becomes less soluble.

As we've shown in this lesson and in previous lessons, polar compounds are able to dissolve polar compounds. The same concept can be applied to non-polar substances, where non-polar compounds are able to dissolve in non-polar substances.



Theory summary

- Water is polar and therefore substances that are polar/charged will be able to dissolve in water.
- The dissolution of ionic compounds does not result in a different substance being produced.
- The solubility of various ionic compounds depends on the types of ions present.
- When substances dissolve in water, the force of attraction holding the substance together needs to be broken, as well as the hydrogen bonding between water molecules.
- A common saying to remember solubility is 'Like dissolves like'.
- The energy produced as a result of interaction between water and the molecular substance/ions in the ionic compound need to be able to overcome the energy required to break the intermolecular bonds/ionic bonds of the substance being dissolved.



7C QUESTIONS

Theory review questions

Question 1

Ionic compounds are all insoluble in water.

A True

B False

Question 2

Ionic substances are able to dissociate in water because

A ion-dipole interactions form between cations and the partially positive charge ends of water molecules.

B ion-dipole interactions form between anions and the partially positive charge ends of water molecules.

Question 3

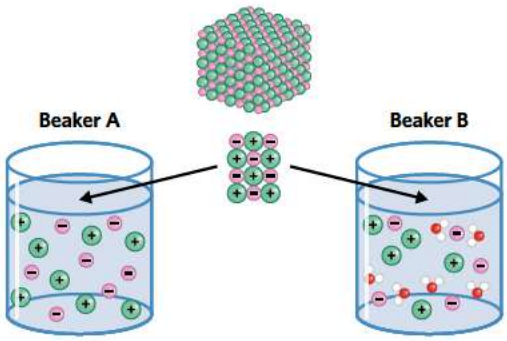
When ionic compounds undergo dissolution, each ion is surrounded by a single water molecule.

A True

B False

Question 4

Which pathway shown in the diagram represents an ionic compound dissolving in water?



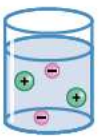
A Beaker A

B Beaker B

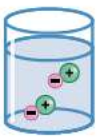
Question 5

When an ionic compound dissociates in water to form a solution, the solution would look like:

A



B



Question 6

Any ionic compound containing sulfate, SO_4^{2-}, is soluble.

A True

B False

Question 7

Which of the following statements is true?

A PbI_2 is soluble.

B $AgNO_3$ is soluble.

Question 8

Non-polar molecular substances are soluble in water.

A True

B False

Question 9

If we were to look under the microscope into a beaker containing water and a soluble alcohol, we would see

A individual alcohol molecules forming hydrogen bonds with water.

B all of the alcohol molecules bonded to each other.

Question 10

All alcohols are soluble in water as they all contain a polar hydroxyl group.

A True

B False

Deconstructed questions

Use the following information to answer questions 11–13.

Potassium iodide, KI, can be used to stop radioactive iodine from being absorbed in the thyroid gland.

Question 11 (1 MARK)

Potassium iodide contains

A K^- ions and I^+ ions.

B Polar bonds.

C K^+ ions and I^- ions.

D metallic bonds.

Question 12 (1 MARK)

When a substance dissolves in water,

A a new product is formed.

B the substance is unchanged.

C it forms interactions with water.

D the atoms are kept in the same arrangement.

Question 13 (2 MARKS)

Draw the interaction that would occur when a sample of potassium iodide is added to water. (2 MARKS)

Exam-style questions

Within lesson

Question 14 (4 MARKS)

Solubility tables are useful when trying to find soluble substances. Using the table as a guide, identify whether each of these salts are soluble or insoluble.



$HgCl_2$



Na_2CO_3



K_3PO_4



$Ca(OH)_2$

Question 15 (4 MARKS)

Magnesium bath salts, one of which is $MgCl_2$, are believed to cause blood vessels to dilate, allowing for increased movement of blood around the body.

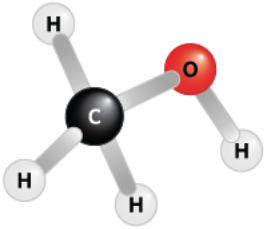
a Write an equation to show the dissociation of $MgCl_2$ in magnesium bath salts. (1 MARK)

b Describe what happens to the chloride ions during dissolution. (3 MARKS)

Question 16 (4 MARKS)

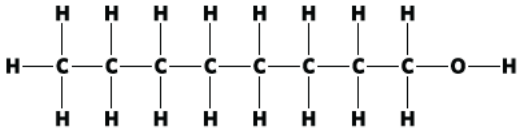
Not all molecular substances can dissolve in water. This depends on the structure of the molecules that make up the substance.

Consider the molecule shown.



a Describe what would happen if a substance consisting of only these molecules was to be added to water. (2 MARKS)

b The molecule shown is part of the same group of molecules as above.



Describe what would happen if a substance consisting of only this molecule was to be added to water. (2 MARKS)

Multiple lessons

Question 17 (5 MARKS)

Epsom salt ($MgSO_4$) is believed to reduce muscle cramping and soreness. Therefore, it is often used by athletes after a competition.

a Write a balanced chemical equation to show the reaction between solutions of epsom salt and sodium chloride (NaCl). (1 MARK)

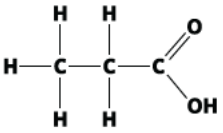
b Calculate the amount, in mol, of epsom salt found in a 200 g sample. (1 MARK)

c Would epsom salt be able to dissolve in water? Explain. (2 MARKS)

d In which group on the periodic table would magnesium be found? (1 MARK)

Question 18 (5 MARKS)

A student wanted to investigate the ability of the molecule shown to dissociate in water.



a Give the name of the molecule shown. (1 MARK)

b Name the functional group present in this molecule. (1 MARK)

c Would this molecule be soluble in water? Explain. (3 MARKS)

Question 19 (4 MARKS)

Methane is a gaseous substance that is often used as a fuel source.

a Draw the structural formula for methane. (1 MARK)

b Can methane dissolve in water? Explain. (2 MARKS)

c Methane is in the same homologous series as the compound with a molecular formula of C_5H_{12}. Give the name of the compound with the molecular formula C_5H_{12}. (1 MARK)

Key science skills

Question 20 (2 MARKS)

Gaya wanted to test the ability of different substances to dissolve in water. To do so, she used the following method:

1. Add 200mL of water into a 500 mL beaker.
2. Add 5.00 g sample of salt.
3. Stir for 2 minutes.
4. Record observations.
5. Repeat for all samples.

Results

a What type of data was collected during this experiment? (1 MARK)

b The results of the experiment are shown. It was known later that the substances tested where: octane, heptane, methanoic acid and magnesium hydroxide.

Substance A	Substance B	Substance C	Substance D
Soluble	Insoluble	Insoluble	Insoluble

Based on the results shown, identify the name of substance A. (1 MARK)

Exam-style question hints

14: The table of solubility outlines the solubility of most ionic compounds. **15a:** Dissociation equations do not include water. **15b:** Chloride ions are anions. **16a:** Molecular substances can be polar. **16b:** The impact of polar covalent bonds diminishes as size of the molecule increases. **17a:** Ionic compounds can take part in double displacement reactions. **17b:** The amount of substance depends on the mass given. **17c:** Not all ionic compounds are soluble. **17d:** Groups refer to columns of the periodic table. **18a:** The names of organic compounds are influenced by the functional groups present. **18b:** Functional groups are groups of atoms that affect the properties of a hydrocarbon. **18c** The solubility of molecular compounds depends on its polarity. **19a:** The structural formula shows all bonds present. **19b:** The ability of molecular compounds to dissolve depends on their polarity. **19c:** Alkanes have an -ane suffix. **20a:** The type of data collected depends on the type of information given. **20b:** Some charged compounds are soluble whereas others aren't.

7D CONCENTRATION

In this lesson, we will learn about the concept of concentration and how the concentration of substances can be expressed in different ways.

7A Properties of water | 7B Thermal properties of water | 7C Dissolution | 7D Concentration | 7E Solubility | 7F Precipitation reactions

Study design dot point

- the concept of solution concentration measured with reference to moles (mol L^{-1}) or with reference to mass or volume (g L^{-1}, mg L^{-1}, %(m/m), %(m/v), %(v/v), ppm, ppb) in selected domestic, environmental, commercial or industrial applications, including unit conversions

Key knowledge units

Concentration	2.2.7.1
Units of concentration of solutions	2.2.7.2
Dilution	2.2.7.3

Key terms and definitions

Concentration amount of chemical substance in a unit of volume

Concentrated solution solution with relatively high amount of solute compared to the volume of solution

Diluted solution solution with relatively low amount of solute compared to the volume of solution

Molarity (M) concentration of a solution measuring the number of moles of solute per litre of solution

Parts per million (ppm) one part in one million parts of a solution

Parts per billion (ppb) one part in one billion parts of a solution

%(m/m) mass percent concentration of solute in solution

%(w/w) weight percent concentration of solute in solution

%(m/v) mass percent concentration of solute per unit of volume of solution

%(v/v) volume percent concentration of solute in solution

Dilute (adj.) solution with decreased concentration of solute

Dilution process of adding solvent to a solution to decrease concentration

Lesson links

This lesson builds on:

- 2F The mole - part 1
- 2G The mole - part 2

Concentration can be calculated based on the amount of substance present.

Concentration 2.2.7.1

OVERVIEW

Concentration measures the amount of solute in a given volume of solution.

THEORY DETAILS

Concentration measures the amount of substance in a given space or volume. For this particular lesson, we will focus on concentration in relation to solutions.

At the most basic level, solutions are made up of a solute and solvent. As shown in figure 1, the solute is the substance being dissolved in the solvent. Therefore, the term concentration refers to the amount of solute present in the solvent. Another way to think about concentration is that it is the ratio of solute to either the solvent or even compared to the whole solution. A solution with lots of solute per unit of volume is described as being a **concentrated solution**, whereas a solution with relatively little solute per unit of volume is referred to as being a **dilute solution** as shown in figure 2.

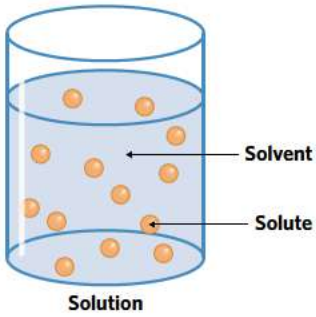


Figure 1 Representation of a solution.

It's important to know that it's not the total number of solute particles that makes a solution concentrated but rather the number of solute particles in a given volume. For the example shown in figure 2, we can see that the concentrated solution contains 12 solute particles for every litre of solution, however the dilute solution only contains three solute particles per litre of solution. As a result, solution A is more concentrated than solution B.

We can quantitatively measure the concentration of a solution in different ways which leads to different units being used to express concentration.

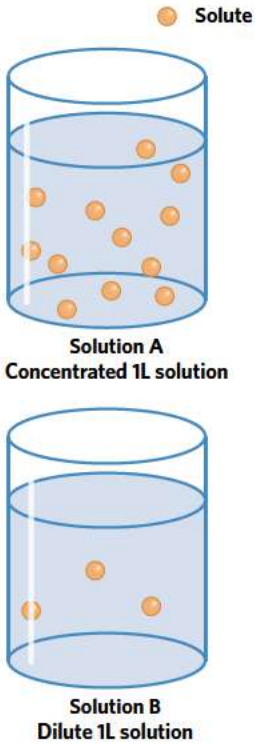


Figure 2 Comparison of a concentrated vs dilute solution.

Units of concentration of solutions 2.2.7.2

OVERVIEW

Concentration can be expressed in different units.

THEORY DETAILS

Molar concentration

The most common way in which we determine the concentration of solutions in chemistry is by calculating the amount, in mol, of solute in a volume of solution. This is known as **molarity**, and is represented by the units mol L^{-1} or M. The volume of the solution is always expressed in litres.

Suppose we added six moles of NaCl solute to one litre of water as shown in figure 3.

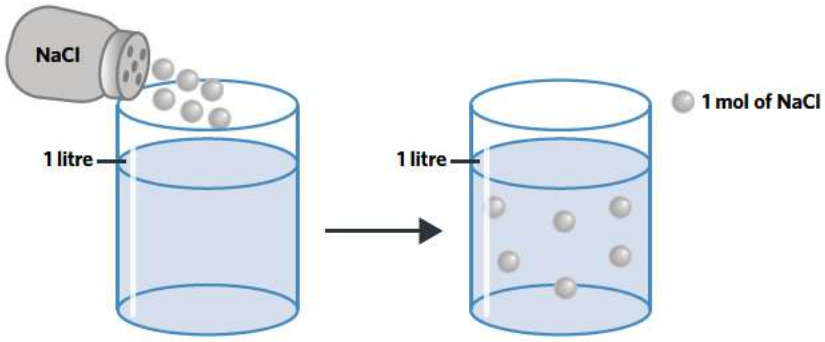


Figure 3 Production of a NaCl solution.

To determine the molarity of this solution, we would do the following:

$$c(\text{NaCl}) = \frac{6 \text{ moles of solute}}{1 \text{ L solution}} = 6 \text{ mol L}^{-1} \text{ (or M)}$$

Therefore, to calculate the molarity of a solution:

Concentration → $c = \frac{n}{V}$ ← Amount / Volume

The amount is expressed in moles (mol) and the volume is measured in litres (L).



For simplicity, we have made each square of NaCl equate to one mole of NaCl, however in reality we know that there are many more particles in a mole of substance and that the NaCl exists as ions in solution.



Always remember to convert the volume of the solution into litres.

Worked example 1

Calculate the molarity of a 100 mL solution containing 2.5 mol of NaCl.

What information is presented in the question?

The volume of the solution, 100 mL and the amount of NaCl, 2.5 mol.

What is the question asking us to do?

Calculate the molarity of the solution.

What strategy(ies) do we need in order to answer the question?

1. Convert the volume into the appropriate units.
2. Calculate the concentration of the solution.

Answer

$$V = \frac{100}{1000} = 0.100 \text{ L}$$

$$c = \frac{2.5}{0.100} = 25 \text{ mol L}^{-1} \text{ or M}$$

We can sometimes be tricked into thinking that solutions with a smaller volume are less concentrated. For example, consider a jug of orange juice that is being poured into two separate cups as shown in figure 4. With no changes being made to either glass, both cups of juice came from the same source, therefore would have the same taste.

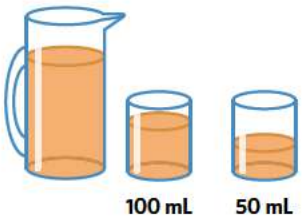


Figure 4 Two glasses of orange juice from the same jug.

At a microscopic view as shown in figure 5, we can see that both cups have different volumes and amounts of solute however they have the same molarity.

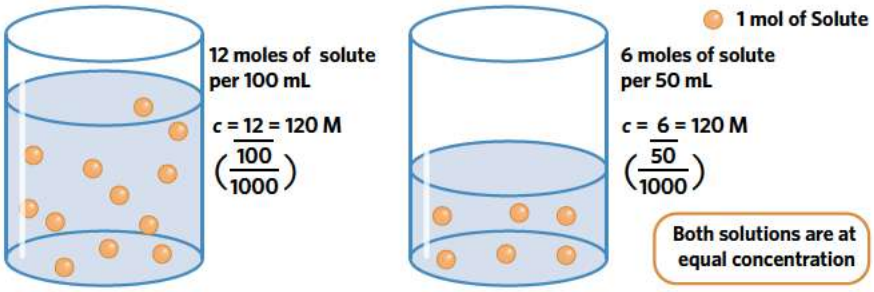


Figure 5 Comparison of solutions of different volumes from the same sample.

We can also use this formula to calculate the concentration, amount of solute or even volume of solutions.



We can use the following concept triangle to help us to figure out the type of calculation to perform to find the concentration, amount of solute or even volume of solutions.

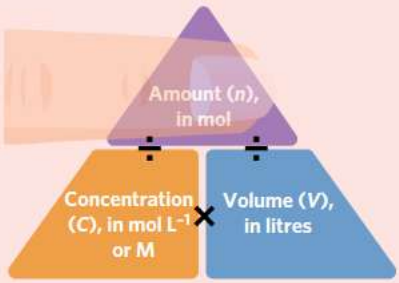


Image: Vasilyeva Larisa/Shutterstock.com

To use this, use your finger to cover up the value you are trying to find. Calculate using the remaining variables as indicated by the signs. For example, to find the amount, cover the section of the triangle that says 'amount', which leaves us to perform the calculation $c \times V$.

As we know from 2F and 2G, the mole concept is connected to many different quantitative values, and can be calculated when given the mass of a substance. We can therefore also calculate molarity when given the mass of solute in a solution.

Worked example 2

5.0 g of HCl was dissolved in 200 mL of water to create a solution. Find the concentration in mol L^{-1} of the HCl solution.

What information is presented in the question?

The volume of the solution, 100 mL and the amount of NaCl, 2.5 mol.

What is the question asking us to do?

Calculate the concentration of the solution, expressed in mol L^{-1}.

What strategy(ies) do we need in order to answer the question?

1. Calculate the amount of HCl present.
2. Convert the volume into the appropriate units.
3. Calculate the concentration of the solution.

Answer

$$n(\text{HCl}) = \frac{5.0}{1.0 + 35.5} = 0.1370 \text{ mol}$$

$$V = \frac{200}{1000} = 0.200 \text{ L}$$

$$c = \frac{0.1370}{0.200} = 0.68 \text{ mol L}^{-1}$$

Worked example 3

A 200 mL sample of a 1.7 M solution of HCl was used in a reaction. What amount, in mol, of HCl is present?

What information is presented in the question?

The volume, 200 mL, and concentration, 1.7 M of HCl.

What is the question asking us to do?

Calculate the number of moles of HCl present.

What strategy(ies) do we need in order to answer the question?

1. Convert the volume into appropriate units.
2. Calculate the amount, in mol, of HCl.

Answer

$V(\text{HCl}) = \frac{200}{1000} = 0.200 \text{ L}$

$c = \frac{n}{V}$ therefore $n = c \times V$

$n(\text{HCl}) = 1.7 \times 0.200 = 0.34 \text{ mol}$

Mass per volume

By law, the products that we purchase from supermarkets must include labelling showing information about the product. Substances such as household cleaners and liquid medication are required to show the concentration of their main ingredients, most of which are expressed as a mass per volume as shown in figure 6. This mainly involves the use of grams per litre, g L^{-1} or milligrams per litre, mg L^{-1} as shown in figure 6.



Figure 6 Example of product showing concentration as mass per volume.

As the description suggests, to calculate concentration based on the mass per volume:

Concentration → $c = \frac{m}{V}$ ← Mass / Volume

Where the volume is expressed in litres (L), and the mass can be given in grams (g) or milligrams (mg).

Worked example 4

Calculate the concentration in mg L^{-1} of a liquid medication that contains 40.0 mg of paracetamol in 100 mL.

What information is presented in the question?

The mass of paracetamol, 40.0 mg, and the volume, 100 mL.

What is the question asking us to do?

Calculate the concentration of the solution, expressed in mg L^{-1}

What strategy(ies) do we need in order to answer the question?

1. Convert the volume into the appropriate units.
2. Calculate the concentration of the solution.

Answer

$V = \frac{100}{1000} = 0.100 \text{ L}$

$c = \frac{40.0}{0.100} = 400 \text{ mg L}^{-1}$

Ppm and ppb

For substances that have very small quantities, we measure concentration in **parts per million (ppm)** or **parts per billion (ppb)**. This refers to a single part of a substance per one million or billion units of another substance as shown in figure 7.

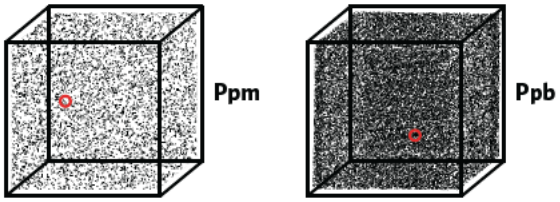


Figure 7 Illustration of ppm and ppb.

These units are used to detect the concentrations of drugs, dangerous metals in food products and drinking water, and even levels of organic pollutants in pesticides. For example, in Victoria, the concentration of fluoride in drinking water cannot exceed 1 ppm; a level recommended by the World Health Organisation (WHO). Fluoride is added into water and toothpastes to help strengthen the structure of teeth.

Ppm and ppb can be expressed using different units so long as what is being compared is 1 millionth or billionth of the substance to which it is being compared to as shown in figure 8 and 9. For example 1 kg consists of 1 000 000 mg. As a result, 1 mg in a 1 kg sample represents 1 ppm.

$$\text{1 kg} = \text{1 000 000 mg} \qquad \frac{\text{Mass of solute (mg)}}{\text{Mass of solute (kg)}} = \text{ppm} = \frac{\text{Mass of solute (mg)}}{\text{Volume of solution (L)}} \qquad \text{1 L (of solution)} = \text{1 000 000 mg}$$

Figure 8 Example of the expression of ppm.

$$\text{1 kg} = \text{1 000 000 000 µg} \qquad \frac{\text{Mass of solute (µg)}}{\text{Mass of solute (kg)}} = \text{ppb} = \frac{\text{Mass of solute (µg)}}{\text{Volume of solution (L)}} \qquad \text{1 L (of solution)} = \text{1 000 000 000 µg}$$

Figure 9 Example of the expression of ppb.

Worked example 5

A solution has the concentration of 1.010 g L^{-1}. Express this concentration as ppm.

What information is presented in the question?

The concentration of a solution, 1.010 g L^{-1}.

What is the question asking us to do?

Convert the concentration of the solution to ppm.

What strategy(ies) do we need in order to answer the question?

1. Determine the reference unit.
2. Determine the unit that is required for ppm.
3. Convert the mass of solute accordingly.

Answer

Based on the concentration given, we need to express our answer 'per L'. To express this concentration to ppm requires the units to change from g L^{-1} to mg L^{-1}.

1.010 g = 1.010 × 1000 = 1010 mg

Therefore, the concentration is 1010 mg L^{-1} or 1010 ppm.

Knowing the units used in different concentration units helps us to be able to convert between them. To do so, we just need to make sure that the quantitative values are expressed in the appropriate units.

Worked example 6

Express the concentration of a 2.0×10^{-9} M $MgCl_2$, solution as ppm.

What information is presented in the question?

The concentration of a solution, 2.0×10^{-9} M.

What is the question asking us to do?

Convert the concentration of the solution to ppm.

What strategy(ies) do we need in order to answer the question?

1. Determine the units required for ppm.
2. Convert any values to the appropriate units.
3. Convert the concentration to ppm.

Answer

The solution is expressed as molarity, which is in mol L^{-1}. Ppm is expressed as mg L^{-1}. Considering both are expressed per litre, we need to convert the mol value to mg.

2.0×10^{-9} M or 2.0×10^{-9} mol L^{-1} means that there are 2.0×10^{-9} moles of solute for every one L.

We need to convert the mole value to mass of $MgCl_2$ in mg.

$m = n \times M$

$m(MgCl_2) = 2.0 \times 10^{-9} \times (24.3 + (2 \times 35.5)) = 1.906 \times 10^{-7}$ g

1.906×10^{-7} g $= 1.906 \times 10^{-7} \times 1000 = 1.91 \times 10^{-4}$ mg

Therefore, the concentration of the solution is 1.9×10^{-4} ppm.

Percentages

Concentration can be expressed as a percentage of solute per solvent. This can be expressed in a number of different ways depending on the state of the solute and solvent. For example, for a solution where both solute and solution quantities are expressed as masses, we can calculate the concentration as % mass/mass (%**(m/m)**) or % weight/weight (%**(w/w)**). Table 1 outlines the common percentage concentrations.

Table 1 Different percentages of concentration.

Concentration	Description	Calculation
% mass/mass %(m/m)	Both solute and solution are expressed as a unit of mass (e.g. mg, g, kg)	$c = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$ * the masses must be in the same units
% weight/weight %(w/w)	Both solute and solution are expressed as a unit of mass (e.g. mg, g, kg)	$c = \frac{\text{weight of solute}}{\text{weight of solution}} \times 100$ * the weights must be in the same units
% mass/volume **%(m/v)**	The solute is expressed in grams whereas the solution is expressed in millilitres (i.e. g/mL)	$c = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100$
% volume/volume **%(v/v)**	Both solute and solution are expressed as a unit of volume (e.g. mL, L)	$c = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$ * the volumes must be in the same units

As these values are expressed as a percentage, the percentage concentration value can be interpreted as an amount per 100. For example, an alcohol with a concentration of 75.0%(v/v) has 75 mL of alcohol in every 100 mL of the beverage, or 75 mL/100 mL. Since the density of water is approximately 1.0 g/mL, we can also convert volumes to masses.

The type of percentage concentration used can vary depending on the substance being analysed.

Worked example 7

A 1.5 L solution contains 21.2 g of NaCl. What is the %(m/v) concentration of the solution?.

What information is presented in the question?

The volume of a solution, 1.5 L and mass of solute, 21.2 g.

What is the question asking us to do?

Convert the concentration of the solution to %(m/v).

What strategy(ies) do we need in order to answer the question?

1. Determine the units used to express %(m/v).
2. Convert the given values to the appropriate units.
3. Calculate the concentration by %(m/v).

Answer

%(m/v) is expressed as g per mL however our values are given as g and L. Therefore, we need to convert the 1.5 L solution to mL.

1.5 L = 1.5 × 1000 mL = 1500 mL

$$\%(m/v) = \frac{21.2}{1500} \times 100 = 1.4\%$$

Dilution 2.2.7.3

OVERVIEW

Diluting a solution makes it less concentrated.

THEORY DETAILS

Earlier in the lesson, we came across the concept of dilute solutions. Since the concentration of a solution measures the amount of solute in a solution, adding more solvent into the solution (without changing the amount of solute) will make the solution more **dilute** as shown in figure 10. This process of decreasing the concentration of a solution is known as **dilution**.

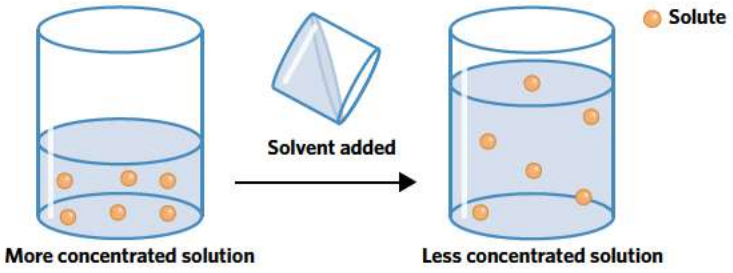


Figure 10 Process of diluting a solution.

As we can see in figure 10, there is no change in the amount of solute in the solution, therefore the amount before and after dilution is the same. Given the molarity of a solution and its volume, the amount of solute is denoted by:

$n = c \times V$

The $n = c \times V$ of the solution before dilution is the same as $n = c \times V$ of the solution after the dilution. This can be expressed as:

Original solution before dilution

$$c_1 \times V_1 = n = c_2 \times V_2$$

Solution after more solvent has been added

Where $c_1 \times V_1$ represents the concentration and volume of solution before dilution, and $c_2 \times V_2$ represents the concentration and volume of solution after dilution. We can simplify this formula to:

Before dilution

$$c_1 \times V_1 = c_2 \times V_2$$

After dilution

Remember that when using this formula, the volume is expressed in litres.



When diluting a solution, the amount (in mol) of solute **does not** change.

Worked example 8

50.0 mL of water was added to a 20.0 mL 2.0 M KCl solution. Calculate the concentration of the diluted solution.

What information is presented in the question?

The volume of water added, 50 mL, and the volume and concentration of the original solution, 20.0 mL and 2.0 M respectively.

What is the question asking us to do?

Calculate the concentration of the diluted solution.

What strategy(ies) do we need in order to answer the question?

1. Determine the final volume of water in the diluted solution.
2. Determine the values for c_1, V_1, c_2, V_2.
3. Calculate the concentration of the diluted solution.

Answer

Since 50 mL of water was added, the final volume of the diluted solution is:

$V_2 = 50.0 + 20.0 = 70.0 \text{ mL}$

$V_{diluted\ solution} = \frac{70.0}{1000} = 0.0700 \text{ L}$

$c_1 = 2.0 \text{ M}$

$V_1 = 20.0 \text{ mL} = \frac{20.0}{1000} = 0.0200 \text{ L}$

$c_2 = ?$

$V_2 = 0.0700 \text{ L}$

$2.0 \times 0.0200 = c_2 \times 0.0700$

$c_2 = \frac{2.0 \times 0.0200}{0.0700} = 0.57 \text{ M}$

Worked example 9

How much water would be required to be added to a 150 mL 2.0 M solution to achieve a final concentration of 1.5 M?

What information is presented in the question?

The volume and concentration of the initial solution, 150 mL and 2.0 M respectively. Also, the desired concentration 1.5 M.

What is the question asking us to do?

Calculate the volume of water required to dilute the solution to 1.5 M.

What strategy(ies) do we need in order to answer the question?

1. Determine the values for c_1, V_1, c_2, V_2.
2. Calculate the final volume of the diluted solution.
3. Calculate the volume of water required.

Answer

$c_1 = 2.0 \text{ M}$

$V_1 = \frac{150}{1000} = 0.150 \text{ L}$

$c_2 = 1.5 \text{ M}$

$V_2 = ?$

$2.0 \times 0.150 = 1.5 \times V_2$

$V_2 = \frac{2.0 \times 0.150}{1.5} = 0.20 \text{ L}$

Volume required = 0.20 − 0.150 = 0.050 L

Theory summary

- Concentration measures the amount of solute per unit of solution.
- Molarity is a common unit of concentration, expressed in mol L^{-1} or M.
- Ppm and ppb measure the amount of solute per million or billion parts of solvent respectively.
- Concentration can also be expressed as %(m/m), %(w/w), %(m/v) and %(v/v).
  - Percentage concentration can also be interpreted as an amount per 100.
- When solutions are diluted, the amount of solute remains the same.

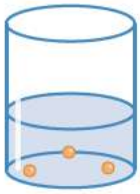
7D QUESTIONS

Theory review questions

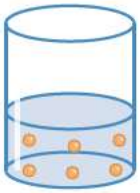
Question 1

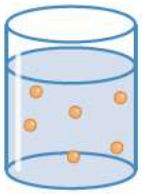
Which of the beakers A or B would have the same concentration as the solution in the reference beaker?

A



B





Reference beaker

Question 2

For molarity, the concentration is expressed as _____ per _____.

A grams, litre

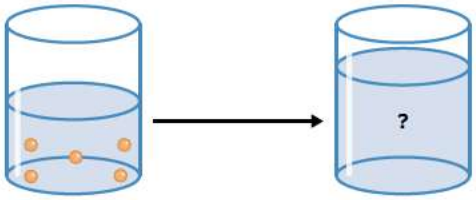
B mol, litre

Question 3

How many particles of solute would be found in the second beaker if both solutions have the same concentration?

A 5

B 10



Question 4

Which of the following concentration units shows concentration at ppm?

A mg/g

B mg/L

Question 5

2.35 grams of $CaCO_3$ was dissolved to make a 30.0 gram solution. Which of the following would be the best way to express the concentration of this solution?

A %(w/w)

B %(m/v)

Question 6

To convert molarity to ppb,

A we would need to convert the amount of solute in mol to µg.

B we would need to convert the volume of solution from litres to µL.

Question 7

To calculate concentration as %(w/w) and %(v/v),

A any unit can be used so long as it fits the unit of measurement.

B the weights and volumes must be in the same units.

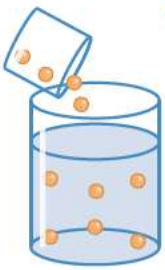
Question 8

Which of the following scenarios illustrates the process of dilution?

A



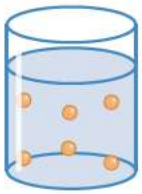
B



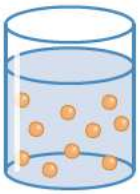
Question 9

Which of the following beakers is more dilute?

A



B



Question 10

When comparing a diluted solution with its original solution,

A they both have the same concentration.

B the both have the same amount of solute.

Deconstructed questions

Use the following information to answer questions 11–13.

A student was asked to make a 500 mL sample of a 2.5 M NaOH solution.

Question 11 (1 MARK)

The solution would contain:

A 2.5 moles of NaOH for every 1.0 L of solution.

B 2.5 g of NaOH for every 500 mL of solution.

C 500 moles of NaOH for every 2.5 g of solution.

D 500 g of NaOH for every 500 g of solution.

Question 12 (1 MARK)

Which of the following formulas would be used to calculate the molarity of the solution?

A $\%(m/m) = \dfrac{\text{mass (in g)}}{\text{mass (in kg)}}$

B $c = \dfrac{\text{amount (in mol)}}{\text{volume (in litres)}}$

C $\text{ppm} = \dfrac{\text{mass (in µg)}}{\text{mass (in kg)}}$

D $\%(v/v) = \dfrac{\text{volume (in mL)}}{\text{volume (in L)}}$

Question 13 (2 MARKS)

Calculate the mass of NaOH found in the solution.

Exam-style questions

Within lesson

Question 14 (23 MARKS)

There are different ways to express concentration.

a For each of the following questions, calculate the concentration in the appropriate units.

- **i** Molarity of a 200 mL solution containing 1.47 g of $CaCl_2$. (2 MARKS)
- **ii** Molarity of a 250 mL solution containing 2.59 g of Na_2CO_3. (2 MARKS)
- **iii** Molarity of 120 mL solution containing 0.966 g of NaCl. (2 MARKS)
- **iv** 200 mL of a NaOH solution containing 2.0 g of NaOH, in ppm. (2 MARKS)
- **v** 200 mL of a NaOH solution containing 2.0 moles of NaOH, in ppm. (2 MARKS)
- **vi** 100 g of solution containing 0.50 mol of $FeSO_4$ in %(w/w). (2 MARKS)

b Convert the units of each solution accordingly.

- **i** 1.0 M of NaCl → ppm (2 MARKS)
- **ii** 1337 ppm → ppb (2 MARKS)
- **iii** 981 ppm KCl → M (2 MARKS)
- **iv** 1.5 M PbI_2 → %(m/v) (2 MARKS)

c Calculate the concentration of the solution for each of the following:

- **i** The final concentration when 25.0 mL of water is added to 20.0 mL of a 1.25 M solution. (1 MARK)
- **ii** The initial concentration of a 20.0 mL sample that was diluted to a 50.0 mL solution of 2.0 M. (1 MARK)
- **iii** The final concentration when 200 mL of water is added to a 1.05 M 150 mL solution. (1 MARK)

Question 15 (3 MARKS)

An iced coffee brand has a nutritional label as shown.

a What is the concentration of sugar in the iced coffee as %(m/v)? (1 MARK)

b Given the sugar has a molar mass of 180 g mol^{-1}, calculate the molarity of sugar in the iced coffee. (2 MARKS)

Nutrition Information	
Servings per pack: 1.5	Serving size: 500 mL
	Ave. QTY. per 100 mL
Protein	3.4 g
Carbohydrate	9.0 g
- Sugars	8.7 g

Question 16 (9 MARKS)

Fish that are sold in Australia cannot have a mercury concentration of more than 1 ppm.

a What is the maximum mass of mercury that can be found in a snapper fish weighing 500 g? (2 MARKS)

b After net fishing in Port Melbourne, Ganesh caught a small fish. All of the mercury was collected and made into a 50 mL solution with a concentration of 1.76 M. What mass of mercury was present in the fish? (2 MARKS)

c Given that the fish had an initial weight of 2.67 kg, would this meet the Australian standards of mercury levels? Explain. (3 MARKS)

d How much water will be required to dilute the mercury solution so that it had a concentration of 1.0 M? (2 MARKS)

Question 17 (5 MARKS)

Sharon bought a bottle of cordial that was shown to have a %(m/v) concentration of 50% worth of sugar ($C_6H_{12}O_6$).

a Calculate the mass of substance in every litre of the cordial solution. (1 MARK)

b Calculate the molarity of the cordial. (2 MARKS)

c Sharon poured a 50 mL sample of the cordial into a cup and wanted to dilute it so that the final sample had a volume of 200 mL. Calculate the concentration of the final solution. (2 MARKS)

Question 18 (6 MARKS)

Pesticides such as aldicarb can leak into waterways and cause serious health risks when consumed by humans. It has been advised that the tolerable concentration of aldicarb in humans is 7.0 ppb a day.

a What mass, in μg, of aldicarb is tolerable in a toddler weighing 13.6 kg? (1 MARK)

b The baby drank 200 mL of water containing 1.82×10^{-5} M aldicarb. Has the baby exceeded her daily limit for aldicarb? Include calculations in your answer. (M_r = 190.3 g mol^{-1} and assume 100% absorption of aldicarb from the drinking water). (5 MARKS)

Multiple lessons

Question 19 (5 MARKS)

Tran wanted to make a 2.5 M KCl solution.

a What type of compound is KCl? (1 MARK)

b Write the equation to show the dissociation of KCl. (1 MARK)

c What mass of solid KCl needs to be added to 100 mL of water to produce the desired solution? (2 MARKS)

d Balance the equation for the reaction between silver nitrate ($AgNO_3$) and potassium chloride. (1 MARK)

Key science skills

Question 20 (5 MARKS)

A student prepared a solution for a reaction by following the method below:

1 Add 250 mL of water into a beaker.

2 Weigh 4.00 g of solid NaCl.

3 Add the NaCl sample into the beaker of water and stir.

a On the way back from the scale, the student tripped over their shoelace and spilled a small amount of NaCl. Since the amount lost was minimal they decided to proceed to make the solution anyway. What type of error is this? (1 MARK)

b How would this affect the concentration of the solution? (2 MARKS)

c What is the theoretical concentration of the solution? (2 MARKS)

Exam-style question hints

14a: Different concentration units use different units of measurement as part of their calculation. **14b:** Unit conversion requires an understanding of the difference in the units used for each concentration unit. **14c:** In a dilution, the amount of solute remains the same. **15a:** The mass and volume need to be expressed in a particular unit of measurement. **15b:** Molarity depends on the amount of solute per unit of volume. **16a:** PPM can be determined by identifying the mass:volume ratio. **16b:** The amount of solute present depends on the volume of the sample. **16c** The mass of substance depends on the amount of substance. **16d:** Dilution changes the concentration but not amount of solute. **17a:** %(m/v) depends on the mass of substance per unit of volume. **17b:** Values presented as a % can be viewed as an amount per 100. **17c:** Dilution changes the concentration but not amount of solute. **18a:** The total mass of aldicarb depends on the total weight. **18b:** To compare data, the values need to be expressed in the same unit of measurement. **19a:** Compounds are generally named based on the type of intramolecular bonding. **19b:** Dissociation equations do not include water. **19c:** Concentration depends on the amount of solute per unit of volume. **19d:** A reaction between ionic compounds results in a 'swapping' of ion pairs. **20a:** Errors are categorised based on the effect on the results. **20b:** Concentration depends on the amount of solute present. **20c:** Concentration depends on the amount of solute present per unit of volume.



7E SOLUBILITY

In this lesson, we will be learning about properties affecting the solubility of substances, the use of solubility curves and how to measure a solution's concentration.

7A Properties of water | 7B Thermal properties of water | 7C Dissolution | 7D Concentration | 7E Solubility | 7F Precipitation reactions

Study design dot points

- the use of solubility tables and experimental measurement of solubility in gram per 100 g of water
- the quantitative relationship between temperature and solubility of a given solid, liquid or gas in water
- the use of solubility curves as a quantitative and predictive tool in selected biological, domestic or industrial contexts

Key knowledge units

Measurement of solubility	2.2.4.2
Effect of temperature on solubility	2.2.5.1
Solubility curves	2.2.6.1

Key terms and definitions

Solute component that dissolves in a solvent

Solubility extent to which a substance dissolves in a solvent

Unsaturated solution solution that can dissolve more solute due to containing less than the maximum amount of solute

Saturated solution solution that cannot dissolve any more solute due to containing the maximum amount of solute

Supersaturated solution solution containing more than the maximum amount of solute that can be dissolved

Crystallisation process by which solute molecules leave solution to form a highly organised structure known as a crystal

Lesson links

This lesson builds on:

▶ 5C Intermolecular bonds

The solubility of a substance depends on the intermolecular interaction between the substance and its solvent.

▶ 7C Dissolution

Substances that can dissolve in another substance are referred to as being soluble.

Measurement of solubility 2.2.4.2

OVERVIEW

Solubility is the degree to which a substance dissolves in a solvent, and varies between concentrations and substances.

THEORY DETAILS

As we have learned previously, dissolution is the process in which a substance is dissolved into another substance. For example, when preparing a drink with powder such as Milo or an energy drink, the powder dissolves in the water it is mixed in making it the **solute**. The water, however, dissolves the powder making it the solvent. As shown in figure 1, solutes dissolve in solvents to form solutions, a process in which the solute becomes dispersed evenly throughout the solution.

In this lesson we will be introducing the concept of **solubility**. While many substances are able to dissolve in solvents, the extent to which they dissolve may differ. For example, electrolytes mix quite easily with water, which means that electrolytes have high solubility in water and are therefore referred to as being water soluble, as shown in figure 2. In contrast, when oil is added to water it will not readily disperse throughout the solution, forming an oil slick instead as shown in figure 2. This is because oil has low solubility in water, and is therefore described as being water insoluble. Rather than dispersing throughout the solvent, the oil clumps together on the surface instead.

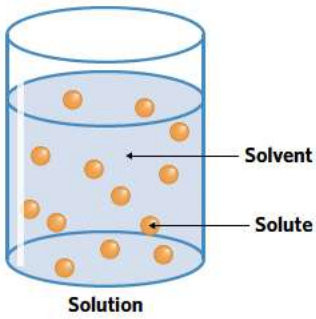
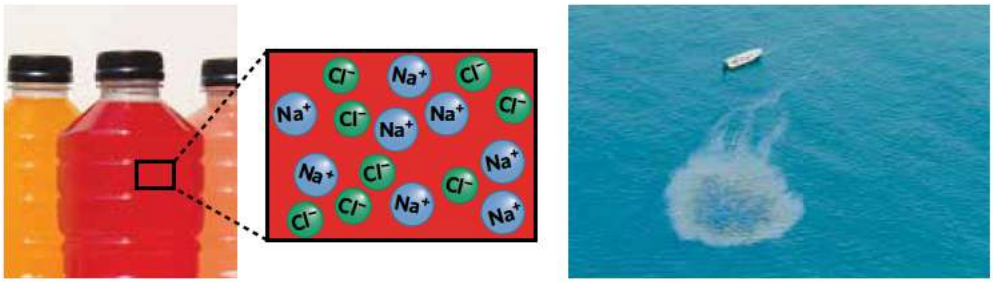


Figure 1 A solution is formed between a solute and a solvent, with the solute evenly distributed throughout it.



Images (left to right): Chuck Wagner, Samjaw/Shutterstock.com

Figure 2 Ions in a sports drink disperse through the water, but oil forms a slick on the surface.

The most common solvent on Earth is by far water, which is sometimes referred to as the 'universal solvent' because of its ability to dissolve a wide array of solutes. However, when using the term solubility, this term simply refers to the extent to which any solute may dissolve in any solvent.

It is also important to note that not all solutes are solids, and that solutes can be in any state. As an example, liquid alcohols can be dissolved in water, thus making it a liquid solute, and carbon dioxide can be a gaseous solute in water. Substances that are dissolved in different states possess distinct properties which will be highlighted later in the lesson.

Types of solutions

It is important to understand that there are three different types of solutions that can be produced through the dissolution of a solute in a solvent. They are classified according to how much solute is dissolved in the solution as described in table 1.

Table 1 The different types of solutions.

Solution type	Description
Unsaturated	A solution containing solute that is not yet saturated and is therefore able to dissolve more solute at a given temperature
Saturated	A solution which cannot dissolve any more solute at a given temperature
Supersaturated	An unstable solution with more dissolved solute than a saturated solution at a given temperature. If disturbed, the solute may form a solid crystal

As we can see in figure 3, the concentration of solute increases from unsaturated solutions to supersaturated solutions.

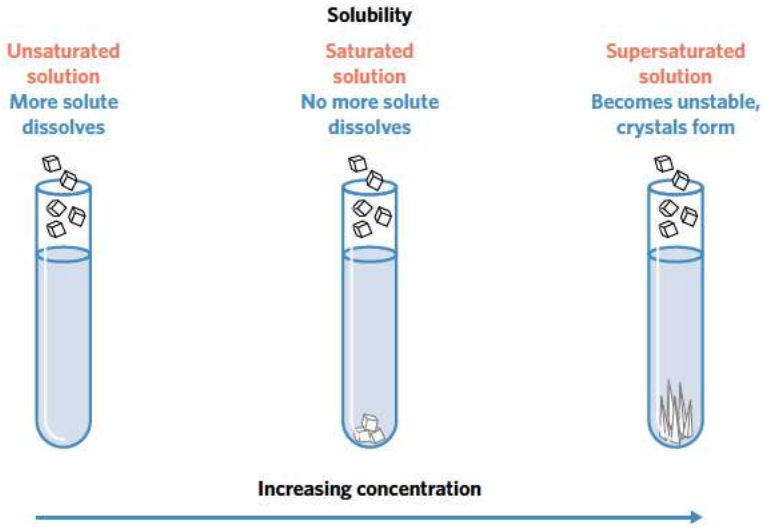


Figure 3 Unsaturated, saturated and supersaturated solutions.

An increase in temperature can increase the solubility of a substance. This is because the increase in energy allows the solvent molecules to be more efficient at breaking the intermolecular bonds holding the solute molecules together. As a result, this allows for more solute to be dissolved in the solvent. We can use this property to produce supersaturated solutions.

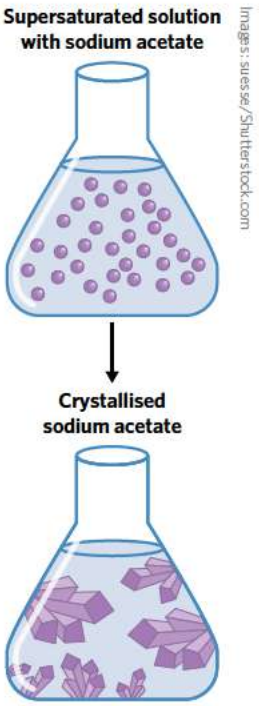
By heating the solution up to a certain temperature, we can increase the solubility of a substance and then cool the solution down to the point where a supersaturated solution is formed. Supersaturated solutions, when disturbed, can form crystals. For example, if we were to shake a supersaturated solution, some solute can separate from the rest of the solution and **crystallise**, as seen in figure 4. The crystals formed have a highly ordered structure and result in a decrease in the amount of solute in the solution to a point where it can cause the solution to transition from a supersaturated to a saturated solution.

As we learned in lesson 7C, a solubility table can be used to determine whether a substance, in particular an ionic substance, is able to dissolve in water. Therefore, we can use the solubility table as shown in table 2 to predict reactions between a solute and water.

Table 2 Solubility table.

Ions	Solubility	Exceptions
Sodium ions: Na^+	Soluble	None
Potassium ions: K^+	Soluble	None
Ammonium salts: NH_4^+	Soluble	None
Chloride ions: Cl^-	Soluble	Insoluble in a compound with Ag^+, Hg^{2+}, Pb^{2+}
Iodide ions: I^-	Soluble	Insoluble in a compound with Ag^+, Hg^{2+}, Pb^{2+}
Nitrates: NO_3^-	Soluble	None
Perchlorates: ClO_4^-	Soluble	None
Sulfates: SO_4^{2-}	Soluble	Insoluble in a compound with Ca^{2+}, Ba^{2+}, Sr^{2+}, Ag^+, Hg^{2+}, Pb^{2+}
Carbonates: CO_3^{2-}	Insoluble	Soluble in a compound with group IA* and NH_4^+
Phosphates: PO_4^{3-}	Insoluble	Soluble in a compound with group IA* and NH_4^+
Hydroxides: OH^-	Insoluble	Soluble in a compound with group IA*, Ca^{2+}, Ba^{2+}, Sr^{2+}
Sulfides: S^{2-}	Insoluble	Soluble in a compound with group IA*, IIA** and NH_4^+

*Group IA metals include Li, Na, K, Rb, Cs, and Fr **Group IIA metals include Be, Mg, Ca, Sr, Ba and Ra



Images: suesse/Shutterstock.com

Figure 4 Formation of crystals from supersaturated solution.



Remember the SNAPE rule in 7C to help you with remembering the solubility of ionic compounds.

Worked example 1

Using a solubility table, determine which of the following ionic compounds will be soluble in water: AgI, Na_2S and KCl.

What information is presented in the question?

The identity of the molecules being analysed.

What is the question asking us to do?

Determine which of the molecules will be soluble in water.

What strategy(ies) do we need in order to answer the question?

1. Find the ions in the compounds in the solubility table.
2. Determine which of the ions are generally soluble in water.
3. Look for any exceptions to solubility.

Answer

AgI: Iodide ions (I^-) are generally soluble in water, however there is an exception when in a compound with silver (Ag^+), and therefore this compound will be water insoluble.

Na_2S: Although sulfide ions (S^{2-}) are generally insoluble, compounds containing sodium (Na^+) are always water soluble.

KCl: Both potassium and chloride are generally water soluble.

Therefore, both Na_2S and KCl are soluble in water.

Effect of Temperature on solubility 2.2.5.1

OVERVIEW

Temperature changes impact the extent to which a solute is able to dissolve into a given solvent.

THEORY DETAILS

While solubility tables provide a general overview of whether an ionic compound will dissolve in water, there are additional factors that contribute to the solubility of a solute. This can be seen in figure 5, which shows how the solubility of sugar increases as the temperature of the solvent increases.

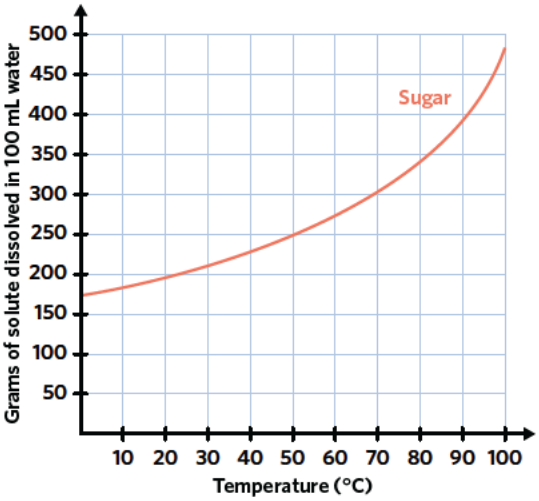


Figure 5 Solubility curve of sugar in water.

As we can see in figure 5, as the temperature of the solution increases, the mass of sugar that can dissolve in 100 mL of water also increases. As a general rule, solubility of a solid solute increases as the temperature increases.

Increasing the temperature does not always enhance solubility and in some cases, increasing the temperature can even decrease the solubility of a solute. For example, increasing the temperature of a cerium (III) sulfate solution reduces the amount of cerium (III) sulfate that can dissolve. However, what is important to understand is that for the majority of cases, there is a direct relationship between temperature and solubility of a solute. This direct relationship between solubility and temperature only applies for solid solutes. When considering the properties of gaseous solutes in a solution, as the temperature of the solution increases the solubility of these substances will decrease. This is shown in figure 6.

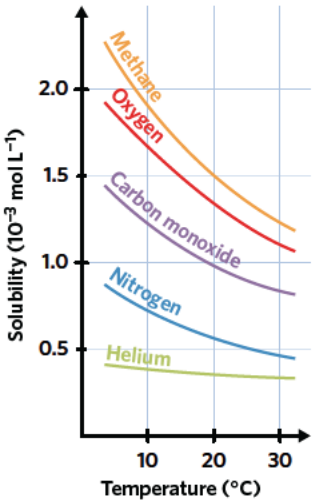


Figure 6 Solubility curve showing the relationship between temperature and the solubility of gases.

This relationship between temperature and solubility is visualised through the construction of a graph known as a solubility curve, which will now be discussed in greater detail.

Solubility curves 2.2.6.1

OVERVIEW

Solubility curves provide a method of visualising and comparing the solubility of substances in a solvent at varying temperatures.

THEORY DETAILS

As we have discovered, different solutes will have different solubilities which can be affected by a change in temperature. The relationship between temperature and solubility can be visualised in the solubility curve shown in figure 7.

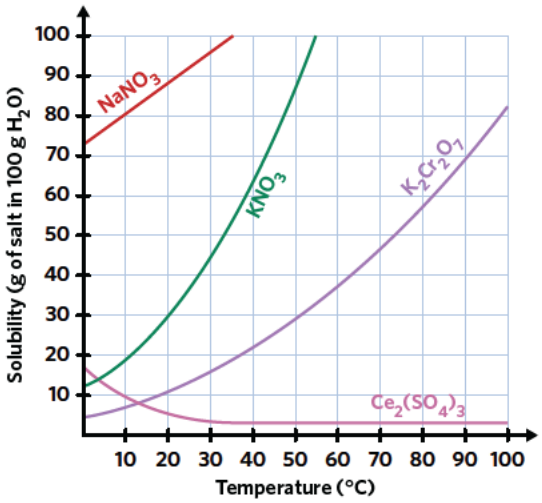


Figure 7 Solubility curves of some common ionic compounds.

When comparing the solubility of different substances, those that have a higher amount of solute dissolved at the same temperature will be more soluble than those with a lower amount of solute dissolved. For example, based on the curves shown in figure 7, we can see that KNO_3 is more soluble than $K_2Cr_2O_7$.

Furthermore, we can see that most curves shown in figure 7 support the idea that an increase in temperature results in an increase in solubility. However, in support of what was mentioned previously, the curve representing $Ce_2(SO_4)_3$ shows the opposite effect, where an increase in temperature results in a decrease in the amount of dissolved $Ce_2(SO_4)_3$ per 100 g of solvent.

Some key features of solubility curves that should be noted include:

- Most solid solutes increase in solubility as temperature increases.
- Each point on the curve represents the maximum amount of solute that can be dissolved in 100 g of water for a given temperature. In other words, each point is the point of saturation for a given temperature.
- The points below the curve represent unsaturated solutions.
- The points above the curve represent supersaturated solutions.

Consider the solubility curve for $K_2Cr_2O_7$ as shown in figure 8.

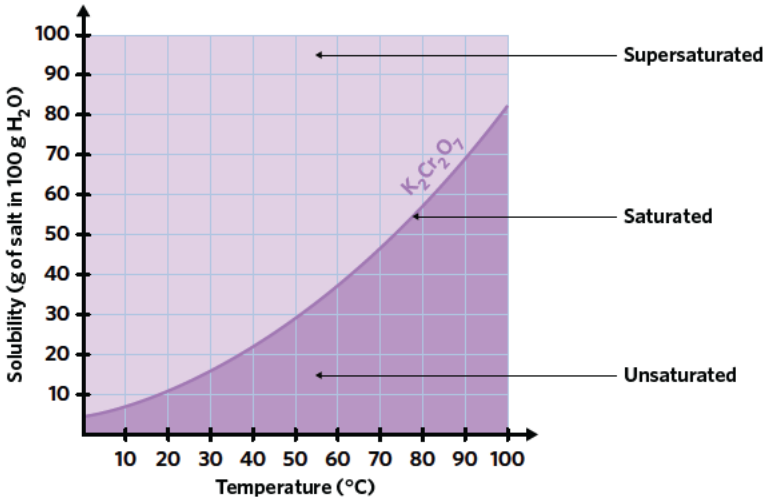


Figure 8 Solubility curve for $K_2Cr_2O_7$.

At 50°C, a saturated $K_2Cr_2O_7$ solution would contain 30 g of solute per 100 mL. Therefore, a solution at 50°C that contains more than 30 g of dissolved solute would be considered supersaturated, whereas a solution at 50°C that contains less than 30 g of dissolved solute would be considered unsaturated. Any undissolved solute in the solution would exist as a solid at the bottom of the solution.

As we can see in the solubility curves presented, the effect of temperature can vary between substances. If the effects of temperature on all solid solutes were the same, it would be unnecessary to construct such a curve. However, since all solutes interact differently with their environments, a solubility curve is necessary to predict how a change in temperature will affect how much of a solute can dissolve in a given solution.

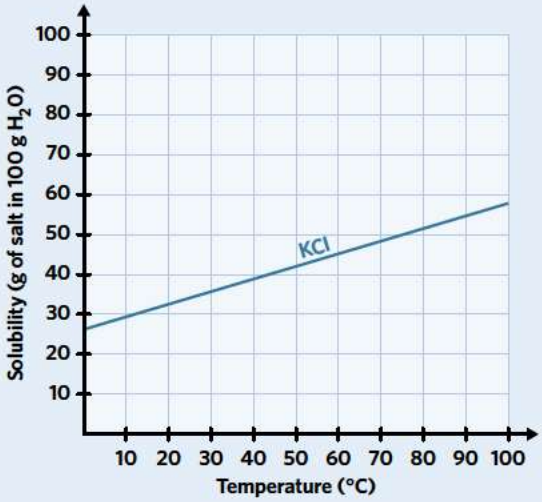
Uses of solubility curves

Solubility curves enable us to visualise and predict the relationship between temperature changes and the solubility of different substances in a solvent. This enables solubility curves to be used in a number of important ways including:

- Predicting the amount of solute that can be dissolved at a given temperature.
- Predicting the amount of solute that won't be dissolved when added to a solvent.
- Predicting the mass of crystals formed in a supersaturated solution.

Worked example 2

Using the following solubility curve, predict the mass, in grams, of KCl that would dissolve in 300 g of H_2O at a temperature of 90°C.



What information is presented in the question?

The identity of the solute, the amount of solvent and the temperature of the solution.

What is the question asking us to do?

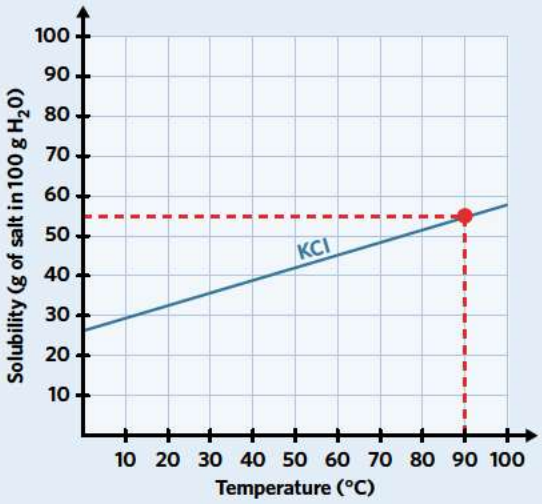
Predict the amount of KCl, in grams, present in 300 grams of water at 90°C.

What strategy(ies) do we need in order to answer the question?

1. Using the curve, find the number of grams of KCl dissolved in 100 g of water at 90°C.
2. Use this value to find the amount of KCl dissolved in 300 g of water.

Answer

According to the solubility curve, 55 grams of KCl can be dissolved in 100 g H_2O at 90°C.



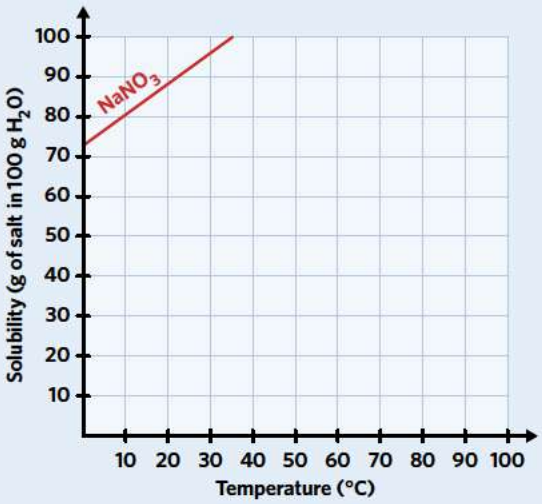
To find the amount of KCl dissolved in 300 g of H_2O, this value must be multiplied by 3.

$55 \times 3 = 165$ g

Therefore, 165 grams of KCl would be dissolved in 300 g of H_2O.

Worked example 3

Using the following solubility curve, predict the mass of crystals produced when 120 g of $NaNO_3$ is added to 100 g of H_2O at 10°C.



What information is presented in the question?

The identity and amount of the solute, the amount of solvent and the temperature.

What is the question asking us to do?

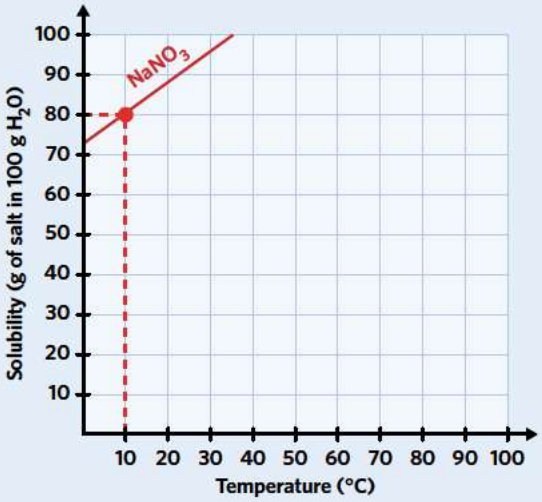
Predict the mass of $NaNO_3$ crystals formed when 120 g of $NaNO_3$ is added to 100 g of H_2O at 10°C.

What strategy(ies) do we need in order to answer the question?

1. Using the curve, find the number of grams of $NaNO_3$ that would dissolve in 100 g of H_2O at 10°C.
2. Find the difference between the amount of solute dissolved in H_2O and the amount added to the solvent to find the degree of supersaturation.
3. Note that the degree of supersaturation relates to the mass of crystals formed.

Answer

According to the solubility curve, 80 g of $NaNO_3$ are dissolved in 100 g H_2O at 10°C.



Crystals are formed through disturbing a supersaturated solution. The amount of crystals formed in grams is the difference between the amount of $NaNO_3$ added to the water and the amount able to dissolve in 100 g H_2O.

120 g – 80 g = 40 g

Therefore, 40 g of $NaNO_3$ crystals will be formed.

As we can see from the examples, by interpreting a solubility curve we can determine the mass of a substance that will dissolve in a quantity of solvent. By doing so, we can also calculate the mass of solute remaining in a solution that can form crystals in supersaturated solutions.

Solubility curves have a number of different domestic and industrial applications. An example of this is in the pharmaceutical industry, where solubility curves are often used to determine the temperature at which a reaction should take place. Additionally, solubility curves are also used in the pharmaceutical industry to determine the point of supersaturation for a solution, as there are some pharmaceuticals that require crystallisation.

Making solubility curves

Solubility curves can be produced through the collection of experimental data. The steps in completing this are detailed below and shown in figure 9:

1 Weigh out several samples ensuring that each sample contains the same mass of solute and record the measurements obtained.
2 Add these samples to a constant volume of water in beakers (or test tubes depending on the volume used). The volume of water must remain constant in all cases.
3 Heat the solutions, stirring the contents until all of the solute has dissolved.
4 Cool the solutions slowly, constantly stirring.
5 Record the temperature at which the crystals from the dissolved solutes begin to form. This is the temperature at which saturation is achieved for the mass of solute present in that test tube.
6 Complete the test for various masses of solute.
7 Plot a solubility curve using the results collected.

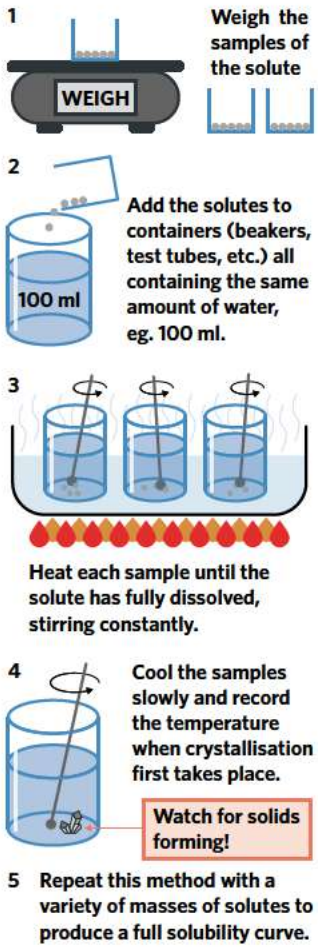


Figure 9 Constructing solubility curves.

Theory summary

- Solubility refers to the extent to which a solute dissolves in a solvent.
- Solutions can be unsaturated, saturated or supersaturated.
- Supersaturated solutions form crystals when disturbed.
- Increasing temperature generally increases solubility of solids and decreases solubility of gases.
- Solubility curves can be used to visualise the relationship between temperature and solubility.

7E QUESTIONS

Theory review questions

Question 1

Which of the following statements is most accurate?

A Increasing temperature increases solubility.
B The effects of increasing temperature on solubility vary from case to case.

Question 2

A saturated solution is

A a solution that can dissolve more solute.
B a solution unable to dissolve more solute.

Question 3

The formation of crystals takes place when

A a supersaturated solution is disturbed.
B a saturated solution is heated.

Question 4

Solubility tables

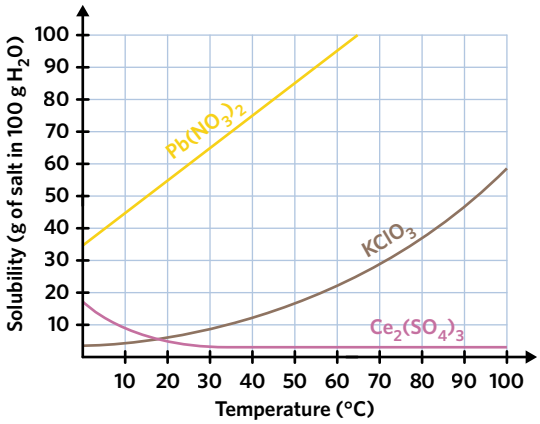
A show the effect of temperature on solubility.
B can be used to predict the solubility of ionic compounds.

Question 5

Using the solubility curve, determine which of the following statements is true:

A Some ionic compound solubilities increase exponentially as temperature increases.

B All ionic compound solubilities increase or decrease linearly against temperature.



Question 6

An unsaturated solution

A is able to dissolve more solute.

B is unable to saturate due to environmental factors.

Question 7

For gaseous solutes, increasing temperatures generally

A increases solubility.

B decreases solubility.

Question 8

The term solubility refers to

A the number of particles dissolved in a solution.

B the extent to which a solute is able to dissolve in a solvent.

Question 9

Concentrations in the area above the line on a solubility curve describe

A saturated solutions.

B supersaturated solutions.

Question 10

Producing a solubility curve involves

A cooling a heated substance and noting when crystals begin to form.

B heating a substance until half of the solute is dissolved.

Deconstructed questions

Use the following information to answer questions 11–13.

Johnson is a chemical engineer looking to find a method to industrialise the large-scale crystallisation of a solute in order to use these for the formation of pharmaceutical solids. To achieve this, Johnson has to know how to use solubility curves to determine the point of crystallisation.

Question 11 (1 MARK)

What is a supersaturated solution?

A An unstable solution with less solute dissolved than a saturated solution.

B A stable solution with more solute dissolved than a saturated solution.

C An unstable solution with more solute dissolved than a saturated solution.

D A stable solution with more solute dissolved than an unsaturated solution.

Question 12 (1 MARK)

How can a solubility curve be used to determine the point of supersaturation?

A A solution is supersaturated when the mass of the solute is above the solubility curve.

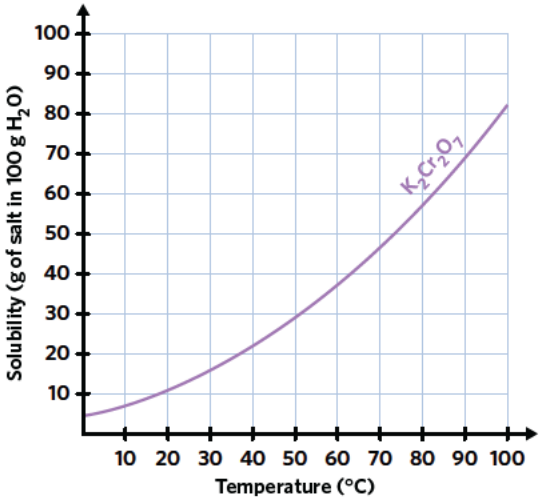
B A solution is supersaturated when a mass of solute is achieved at a temperature higher than that shown on the solubility curve.

C A solution is supersaturated when the mass of the solute is higher than 100 g/100 g H_2O.

D Any solubility above the initial point on a solubility curve is considered supersaturated.

Question 13 (3 MARKS)

Using the given solubility curve, explain if and why crystals will be formed if 180 g of $K_2Cr_2O_7$ is dissolved in 200 g of H_2O at 90°C. If crystals are formed, calculate the mass of the crystals formed.



Exam-style questions

Within lesson

Question 14 (6 MARKS)

Solubility curves are useful for visualising the relationship between temperature and solubility in the dissolution process.

a Use the data given to draw a graph to plot the saturation points at different temperatures. (2 MARKS)

- 10 g per 100 g H_2O at 20°C
- 25 g per 100 g H_2O at 40°C
- 40 g per 100 g H_2O at 50°C
- 100 g per 100 g H_2O at 60°C

b Using the plotted points, draw the solubility curve of this solute. (2 MARKS)

c Comment on the shape of the solubility curve. What implications does this have regarding the relationship between increasing temperature and solubility? (2 MARKS)

Question 15 (8 MARKS)

Solubility tables provide information about ionic compounds, such as which anions and cations are soluble or insoluble in water, and which exceptions exist. Emily, a year 11 chemistry student, has become interested in using a solubility table in order to predict which of her ionic compounds dissolve in water.

a Using the solubility table in table 2, predict which of the following compounds are soluble in water: (2 MARKS)

i KCl **ii** HgS **iii** $CaSO_4$ **iv** Na_3PO_4

Using the information collected in part a, Emily sets about dissolving the soluble compounds in water, in order to observe how they interact with water. She notices that solids begin to appear some time after adding solute to the solution.

b Emily claims that these solutes must be insoluble in water due to the presence of these solids. Is Emily correct? Explain your reasoning. (3 MARKS)

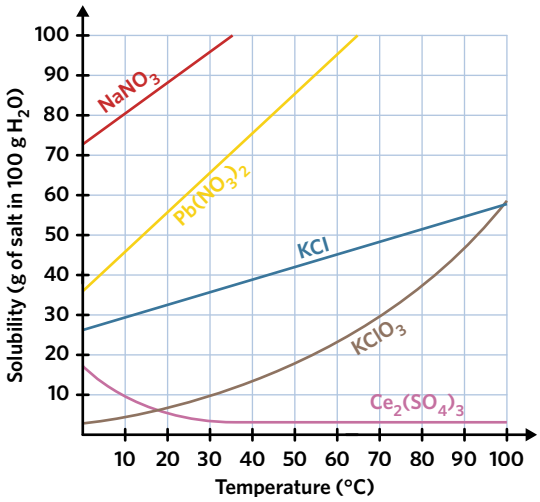
c What would happen if Emily added an insoluble compound to the water? How do the interactions of the insoluble compound in water differ from the interactions of a substance which forms crystals in water? Justify your answer. (3 MARKS)

Question 16 (8 MARKS)

A class is using a solubility curve to determine the relationship between temperature and solubility of ionic compounds.

a Identify the most common relationship between temperature and solubility of ionic compounds. (1 MARK)

b Using the solubility curve provided, identify any compounds that do not follow the most common relationship between temperature and solubility. Determine the relationship found in this case. (2 MARKS)



c Using the solubility curve provided, determine whether the following solutions are unsaturated, saturated or supersaturated:

- **i** 40 g KCl in 100 g H_2O at 70°C. (1 MARK)
- **ii** 10 g $KClO_3$ in 100 g H_2O at 30°C. (1 MARK)
- **iii** 70 g $Pb(NO_3)_2$ in 100 g H_2O at 34°C. (1 MARK)
- **iv** 80 g $NaNO_3$ in 95 g H_2O at 10°C. (1 MARK)
- **v** 25 g $Ce_2(SO_4)_3$ in 450 g H_2O at 20°C. (1 MARK)

Question 17 (7 MARKS)

A year 11 chemistry student has recently learned about solubility, and wishes to test the theory through conducting a series of experiments. She sets up a series of test tubes, each containing solute in a solution with water. The student then places the test tubes in a water bath that is exposed to a heat source.

a The student wishes to determine the temperature at which a particular amount of solute becomes a saturated solution. Describe the experimental technique that should be used. (4 MARKS)

b The student notes that a particular solution is saturated with 55 g of solute in 100 g of H_2O at a temperature of 50°C. Determine whether the solution will be unsaturated, saturated or supersaturated in the following scenarios, or whether the answer cannot be determined:

- **i** The amount of H_2O is increased to 150 g. (1 MARK)
- **ii** The amount of solute is increased to 65 g. (1 MARK)
- **iii** The temperature is increased to 60°C. (1 MARK)

Multiple lessons

Question 18 (16 MARKS)

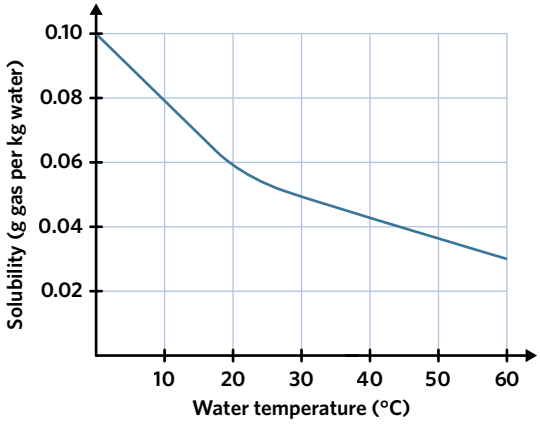
Dissolution is the process in which a substance gets dissolved in another substance. When referring to the dissolution process, a number of important terms are used.

a Define the following terms:

- **i** solute (1 MARK)
- **ii** solvent (1 MARK)

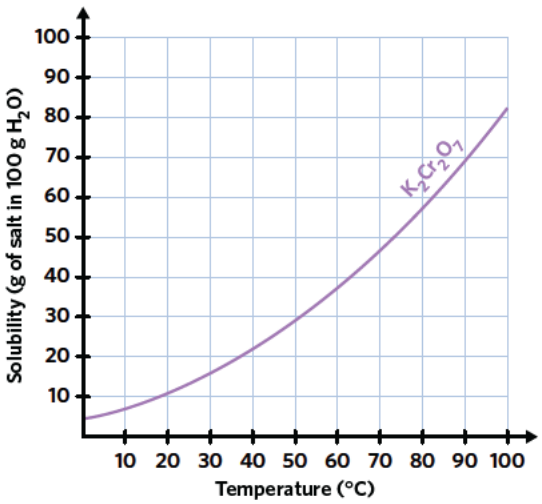
b Water is sometimes referred to as the universal solvent, as it is able to dissolve such a large variety of substances. Due to the structure of water molecules, water is far better at dissolving polar molecules than non-polar substances. Will ethanol or ethane have a higher solubility in water? Justify your answer with reference to the structures of each molecule. (3 MARKS)

c Ethane is a hydrocarbon with a boiling point of −89°C. The solubility curve of ethane in a solution of water is shown below:



- **i** Does the solubility curve for ethane follow the expected relationship between temperature and solubility? Explain your answer. (3 MARKS)
- **ii** Ethanol has a boiling point of 78.4°C. What differences would you expect in the solubility curve for ethanol when compared to the solubility curve of ethane? Explain whether you can predict these differences with certainty. (3 MARKS)

d $K_2Cr_2O_7$ is a substance whose solubility increases as the temperature increases, as can be seen in the graph.



i Using the graph, determine the mass of $K_2Cr_2O_7$ that will be dissolved in 500 g H_2O at 50°C. (2 MARKS)

ii Calculate the amount, in mol, of $K_2Cr_2O_7$ that will be dissolved in 500 g H_2O at 50°C. (1 MARK)

iii If the density of water is 1 g/mL, determine the concentration of a saturated solution of of $K_2Cr_2O_7$ in 500 g of H_2O at 50°C. (2 MARKS)

Question 19 (6 MARKS)

Riley is a chemistry student attempting to produce her own solubility table through a thorough analysis of a number of different chemical experiments. In order to determine whether a solute is soluble in water, she adds the solute to a set volume of water, leaves the solution unobserved for ten minutes to allow it to dissolve and then, holding the solution up to a light, determines whether there are solid particles present. She then collates all her results and organises them in a solubility table.

a Is this likely to be an effective method of establishing a solubility table? Explain. (3 MARKS)

b Having discovered that her current process of producing a solubility table was not working, Riley uses some of her prior chemistry knowledge to gain information for her solubility table.

i Using the periodic table, list the names and chemical symbols of the elements that are in group IA. (2 MARKS)

ii Identify at least two elements in group IA of the periodic table that are soluble in water. (1 MARK)

Key science skills

Question 20 (9 MARKS)

Luke is a year 11 chemistry student who is attempting to construct a solubility curve. In constructing the curve, Luke uses a series of different soluble ionic compounds, and compares their solubilities at different temperatures to produce a curve. He does this through measuring the mass of a solute, dissolving it and heating it in water, and subsequently cooling it and noting the temperature when crystallisation takes place. However, when plotting the curve, Luke notices that some of the values differ greatly from the continuous exponential curve produced by the other values.

a What type of error has Luke made here? (1 MARK)

b Provide two examples of an error Luke may have made to produce these results. (2 MARKS)

After repeating the experiment, Luke noticed that all of his results were centred around the expected value, however the individual results were all different to each other.

c Comment on the accuracy and precision of Luke's experimental results. (2 MARKS)

d How could Luke improve the results of the experiment? Provide an example of how to achieve this. (2 MARKS)

e Why would it be important to stir the solution while it is being heated and while it is being cooled? (2 MARKS)

Exam-style question hints

14a: The mass of solute dissolved is expressed on the y-axis. **14b:** The solubility curve is a continuous curve connecting the plotted points. **14c:** The effect of temperature on the solubility of the solute changes. **15a:** Different ions have varying levels of solubility. **15b:** Solutions can be unsaturated, saturated or supersaturated. **15c:** Crystallisation can occur for soluble substances. **16a:** Graphs show the relationship between the independent and dependent variables. **16b:** Not all graphs show the same relationship. **16c:** Unsaturated solutions contain masses below the curve, whereas supersaturated solutions contain masses above the curve. **17a:** Crystals form as solutions transition from supersaturated to saturated. **17b:** Supersaturated solutions contain more solute than saturated solutions. **18a:** Solute and solvent together make up a solution. **18b:** Ethanol molecules possess a hydroxyl functional group. **18c:** Solubility depends on the strength of intermolecular bonds. **18d:** The solubility curve gives information about the amount of solute that can be dissolved at different temperatures. **19a:** The final result of an interaction does not convey all of the information about the interaction. **19bi:** The periodic table can be divided into groups and periods. **19bii:** Consider the SNAPE rule. **20a:** If some values are greatly off the expected value, then only they are being impacted by the error. **20b:** Random changes to any variable in the experiment can cause this type of error. **20c:** Accuracy and precision describe two different properties of experimental results. **20d:** Imprecision comes about as a result of variations in technique. **20e:** Consistency is important when assessing the composition of a substance.

7F PRECIPITATION REACTIONS

In this lesson, we will be learning about the formation of solids called precipitates and how to write chemical equations for these reactions.

7A Properties of water | 7B Thermal properties of water | 7C Dissolution | 7D Concentration | 7E Solubility | 7F Precipitation reactions

Study design dot point

- precipitation reactions represented by balanced full and ionic equations, including states

Key knowledge units

Formation of a precipitate	2.1.4.1
Writing equations for precipitation reactions	2.1.4.2

Key terms and definitions

Precipitation reaction chemical reaction which results in the formation of a solid from an aqueous solution

Precipitate solid compound formed in a precipitation reaction

Full equation chemical equation with all aqueous ions and chemical species present

Spectator ions aqueous ions which do not react and remain dissolved in solution throughout a reaction

Net ionic equation full equation once the spectator ions have been removed showing only the chemical species that have taken part in a reaction

Lesson links

This lesson builds on:

- 7E Solubility

The solubility of a substance is important in determining whether a precipitate forms.

Formation of a precipitate 2.1.4.1

OVERVIEW

Some chemical reactions produce a solid precipitate in what is known as a precipitation reaction.

THEORY DETAILS

In a chemical reaction, reactants undergo changes to their structure and bonding, resulting in the formation of products. These products may have different chemical and physical properties to the reactants since they are different compounds. **Precipitation reactions** occur when two aqueous solutions react to form an insoluble product. Since the product is insoluble, it exists as a solid in the solution and is known as a **precipitate**. Whether a precipitate forms is based on the solubility rules learnt in lesson 7E.

Precipitation reactions are very useful in the world of chemistry. This is because solid precipitate formed in an aqueous solution can be easily removed from the solution whereas aqueous ions cannot be. This is opposed to trying to remove aqueous ions in a solution which is almost impossible. One of the applications of this is gravimetric analysis which will be discussed later. For example, precipitation reactions are used to remove metal contaminants from water as shown in figure 1.

Let's look at an example of a precipitation reaction. According to the solubility rules, sodium chloride (NaCl) is soluble in water as compounds with Na^+ ions are soluble in water. Silver nitrate ($AgNO_3$) is also soluble in water according to the solubility rules, as compounds with NO_3^- ions are soluble in water. Therefore, if we place solid NaCl into one beaker of water, and solid $AgNO_3$ into another beaker of water, both ionic compounds will dissociate into the ions that they consist of, as shown in figure 2.



Image: Lindsey Moore/Shutterstock.com

Figure 1 Formation of a yellow precipitate (lead (II) iodide).

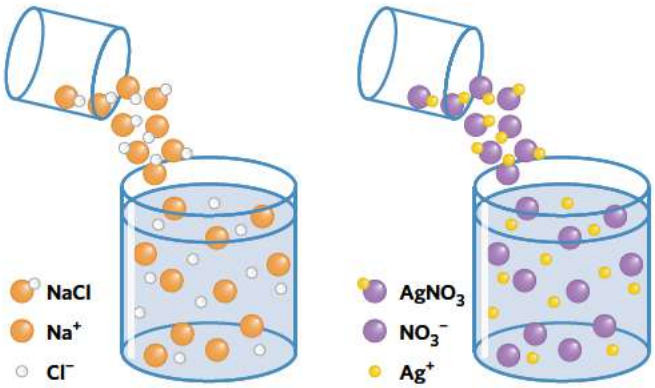


Figure 2 Beakers with solutions of $NaCl_{(aq)}$ and $AgNO_{3(aq)}$.

When these two solutions are mixed together, a solution of $Na^+_{(aq)}$, $Cl^-_{(aq)}$, $Ag^+_{(aq)}$ and $NO_3^-{}_{(aq)}$ is created as shown in figure 3. In this new solution, all four different types of ions are freely moving around and bumping into each other. This means that new combinations of positively and negatively charged ions can be created. For example, $Na^+_{(aq)}$ could form an ionic bond with not only its original partner $Cl^-_{(aq)}$ but also a new partner $NO_3^-{}_{(aq)}$. In a similar way, $Ag^+_{(aq)}$ could form an ionic bond with its original partner $NO_3^-{}_{(aq)}$ as well as a new partner $Cl^-_{(aq)}$. Therefore, there are four possible ionic compounds that could form as a result of mixing these two solutions together, as shown in figure 4. If all of the ionic compounds that form are soluble, all species will remain as dissolved ions and no chemical reaction will take place. If one (or more) of the ionic compounds is insoluble, however, a solid insoluble ionic compound will be produced. This solid is known as the precipitate.

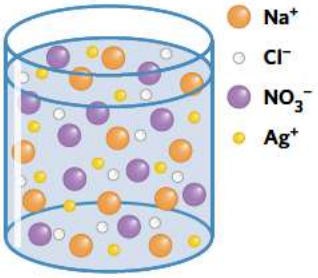


Figure 3 A beaker with $NaCl_{(aq)}$ and $AgNO_{3(aq)}$ solutions.



Figure 4 Possible combinations of ionic compounds.

Using the example above, $NaNO_{3(aq)}$ is soluble in water because compounds with Na^+ ions are soluble in water according to the rules of solubility. $AgCl_{(s)}$ however, is insoluble in water because while $Cl^-_{(aq)}$ tends to form soluble compounds according to the rules of solubility, the combination of $Ag^+_{(aq)}$ ions with $Cl^-_{(aq)}$ ions is an exception. Therefore, a precipitation reaction will occur as shown by the following equation:

$$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow NaNO_{3(aq)} + \underset{\text{Solid precipitate}}{AgCl_{(s)}}$$

Figure 5 shows the reaction represented by this equation.

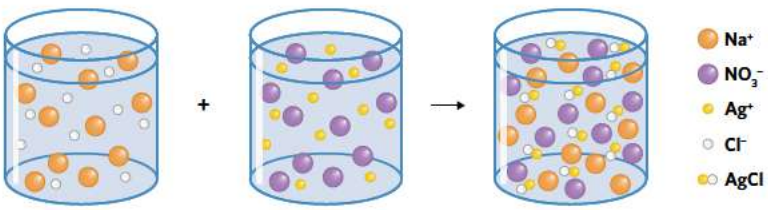


Figure 5 The precipitation reaction between $NaCl_{(aq)}$ and $AgNO_{3(aq)}$.

During a chemical reaction, the precipitate can sometimes be seen as a solid that settles to the bottom of the beaker or test tube. Other times however, the precipitate is small, resulting in all of the particles dispersing throughout the solution which makes the solution look cloudy.



Use the solubility table to help you to determine the precipitate (insoluble substance) in the reaction.

Worked example 1

A student is investigating the properties of ionic compounds and mixes together two aqueous solutions of $CaCl_{2(aq)}$ and $Na_2CO_{3(aq)}$. The solution suddenly becomes cloudy. With reference to precipitation reactions, explain this observation.

What information is presented in the question?

$CaCl_{2(aq)}$ and $Na_2CO_{3(aq)}$ are the reactants.
The solution turns cloudy.

What is the question asking us to do?

Use the chemical theory of precipitation reactions to explain the observation.

What strategy(ies) do we need in order to answer the question?

1. Find out what ions are dissolved in solution.
2. Find out what ionic compounds are able to be formed between positively and negatively charged ions.
3. Use the solubility rules to determine which ionic compounds are soluble or insoluble.
4. Based on this, explain the observation of the solution turning cloudy.

Answer

When solutions of $CaCl_{2(aq)}$ and $Na_2CO_{3(aq)}$ are mixed together, there will be $Ca^{2+}_{(aq)}$, $Cl^-_{(aq)}$, $Na^+_{(aq)}$ and $CO_3^{2-}{}_{(aq)}$ ions in solution. Therefore, $CaCO_3$ and NaCl are the new ionic compounds that are able to be formed. According to the solubility rules, NaCl is soluble whereas $CaCO_3$ is insoluble. This means that $CaCO_3$ will be a precipitate. Since it is solid, rather than dissolved in solution, it will turn the solution cloudy.

Writing equations for precipitation reactions 2.1.4.2

OVERVIEW

Precipitation reactions can be expressed using equations.

THEORY DETAILS

Like all other chemical reactions, precipitation reactions can be shown using chemical equations. The key feature of a precipitation reaction is that a solid is formed. This is expressed with the (s) subscript as shown in the following equation. It is important to include all the states associated with each chemical species because precipitation reactions are defined by the precipitation, or creation of a solid.

Consider a reaction between calcium chloride ($CaCl_2$) and sodium carbonate (Na_2CO_3) solutions. This reaction results in the formation of a precipitate calcium carbonate ($CaCO_3$) and soluble sodium chloride (NaCl). This chemical reaction would be represented as:

$$CaCl_{2(aq)} + Na_2CO_{3(aq)} \rightarrow CaCO_{3(s)} + 2NaCl_{(aq)}$$

This is also known as the **full equation** because all the chemical species are present, even those that have remained unchanged throughout the reaction. In this case, $Na^+_{(aq)}$ and $Cl^-_{(aq)}$ exist in solution before and after the reaction. Since they are not involved in the precipitation reaction, they are referred to as **spectator ions**. They are simply spectators of the main precipitation reaction as shown in figure 6. This means they exist in the aqueous state on both sides of the reaction and have not undergone a chemical reaction.

If the spectator ions, $Na^+_{(aq)}$ and $Cl^-_{(aq)}$ are removed from the full equation, the **net ionic equation** is generated. The net ionic equation is the full equation, once the spectator ions have been removed, and represents only the chemical species which are reacting to form the solid precipitate as shown in figure 7.

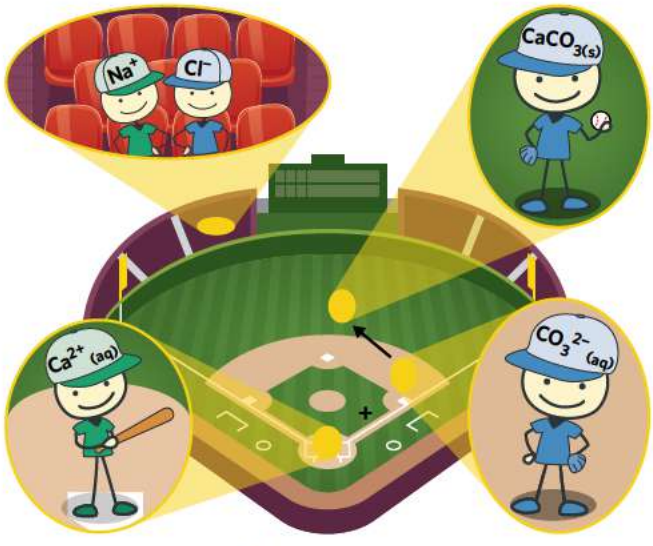


Image: kintomo, klyaksun/Shutterstock.com

Figure 6 Spectator ion analogy.

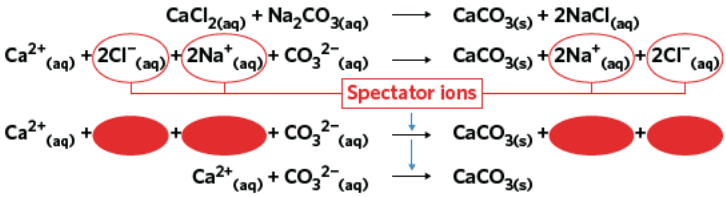


Figure 7 Generation of a net ionic equation.

Another way to generate the ionic equation for a precipitation reaction is to look just at the precipitate that is formed. For example, in one precipitation reaction, the precipitate barium sulphate, $BaSO_{4(s)}$, is produced. Therefore, we know that on the right hand side of the ionic equation, barium sulfate will be the only product.

$\rightarrow BaSO_{4(s)}$

The ions that make up $BaSO_{4(s)}$ are Ba^{2+} and SO_4^{2-}. These will react in a 1 : 1 ratio to form the solid precipitate $BaSO_4$. Since this is a precipitation reaction where a solid is produced from an aqueous solution, Ba^{2+} and SO_4^{2-} are aqueous ions. The formation of solid $BaSO_4$ can be represented in the following ionic equation:

$Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \rightarrow BaSO_{4(s)}$

Worked example 2

Sodium hydroxide, $NaOH_{(aq)}$, reacts with copper (II) sulphate, $CuSO_{4(aq)}$, to produce a pale blue precipitate. Identify the precipitate and write the full equation, as well as the ionic equation.

What information is presented in the question?

The reactants are $NaOH_{(aq)}$ and $CuSO_{4(aq)}$

A precipitate is formed.

What is the question asking us to do?

Determine the chemical formula for the precipitate.

Write out the full chemical equation.

Write out the ionic equation.

What strategy(ies) do we need in order to answer the question?

1. Determine the aqueous ions that are initially present in the solution.
2. Determine the new ionic compounds that can form from the mixing of positively and negatively charged ions together.
3. Use the solubility rules to determine which ionic compound is insoluble and therefore the precipitate.
4. Write out the full equation.
5. Identify the spectator ions.
6. Write out the ionic equation which is the full equation without the spectator ions.

Answer

When solutions of $NaOH_{(aq)}$ and $CuSO_{4(aq)}$ are mixed together, there will be $Na^+_{(aq)}$, $OH^-_{(aq)}$, $Cu^{2+}_{(aq)}$ and $SO_4^{2-}_{(aq)}$ in solution. Therefore, Na_2SO_4 and $Cu(OH)_2$ are the new ionic compounds that will form. According to the solubility rules, Na_2SO_4 is soluble whereas $Cu(OH)_2$ is insoluble. This means that $Cu(OH)_{2(s)}$ is the precipitate.

The full equation, including states, is therefore:

$2NaOH_{(aq)} + CuSO_{4(aq)} \rightarrow Cu(OH)_{2(s)} + Na_2SO_{4(aq)}$

Since $Na^+_{(aq)}$ and $SO_4^{2-}_{(aq)}$ appear on both sides of the equation as aqueous ions, they are spectator ions. Therefore, the ionic equation is the full equation without spectator ions as shown.

$2OH^-_{(aq)} + Cu^{2+}_{(aq)} \rightarrow Cu(OH)_{2(s)}$

Theory summary

- Precipitation reactions involve the formation of a solid from an aqueous solution.
- Precipitation reactions occur due to different solubilities of ionic compounds.
- The precipitate is the solid formed in the reaction and is denoted by the (s) subscript.
- The full equation shows all the ions present, including spectator ions.
- Spectator ions are not involved in precipitation reactions and remain dissolved in solution throughout the reaction.
- The ionic equation is the full equation, with the spectator ions removed.

7F QUESTIONS

Theory review questions

Question 1

Precipitation reactions convert solids into aqueous solutions.

A True

B False

Question 2

In the following image, two clear solutions are mixed together.

The image shows the

A dissolution of a solid.

B formation of a precipitate.



Question 3

Which of the following chemical equations represents a precipitation reaction?

A $Pb(NO_3)_{2(aq)} + 2NaI_{(aq)} \rightarrow PbI_{2(s)} + 2NaNO_{3(aq)}$

B $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

Use the following information to answer questions 4–7.

$BaCl_{2(aq)}$ and $Na_2SO_{4(aq)}$ are mixed together in a test tube.

Question 4

Which option correctly identifies the precipitate formed?

A $BaSO_4$

B NaCl

Question 5

What is the full equation?

A $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$

B $Ba^{2+}_{(aq)} + SO_4{}^{2-}_{(aq)} \rightarrow BaSO_{4(s)}$

Question 6

What is the ionic equation?

A $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$

B $Ba^{2+}_{(aq)} + SO_4{}^{2-}_{(aq)} \rightarrow BaSO_{4(s)}$

Question 7

What are the spectator ions?

A $Ba^{2+}_{(aq)}$, $SO_4{}^{2-}_{(aq)}$

B $Na^{+}_{(aq)}$, $Cl^{-}_{(aq)}$

Question 8

Precipitation reactions are useful because

A they can remove unwanted contaminants from a water sample.

B they don't depend on the solubility of the ionic compounds present.

Question 9

An industrial chemist wanted to remove the valuable Ag^+ ions from solution. Which one of the following solutions would produce a precipitate when mixed with a solution containing Ag^+ ions?

A $NaCl_{(aq)}$
B $NaNO_{3(aq)}$

Question 10

What is the ionic equation when sodium hydroxide, $NaOH_{(aq)}$ and iron (II) sulphate, $FeSO_{4(aq)}$ are mixed together?

A $Na^+_{(aq)} + SO_4{}^{2-}_{(aq)} \rightarrow Fe(OH)_{2(s)} + Na_2SO_{4(aq)}$
B $Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2(s)}$

Deconstructed questions

Use the following information to answer questions 11–13.

Eloise a chemistry student is investigating the properties of different aqueous solutions and whether or not they form precipitates when mixed together. At her disposal are solutions of $Co(NO_3)_{2(aq)}$, $NaOH_{(aq)}$, $KNO_{3(aq)}$ and $CaCl_{2(aq)}$.

Question 11 (1 MARK)

Which two solutions will form a precipitate when mixed together?

A $Co(NO_3)_{2(aq)}$ and $NaOH_{(aq)}$
B $Co(NO_3)_{2(aq)}$ and $KNO_{3(aq)}$
C $Co(NO_3)_{2(aq)}$ and $CaCl_{2(aq)}$
D $KNO_{3(aq)}$ and $CaCl_{2(aq)}$

Question 12 (1 MARK)

Which of the following statements is incorrect regarding spectator ions?

A Spectator ions are in the aqueous state on both sides of the chemical equation.
B Spectator ions are not involved in precipitation reactions.
C Spectator ion identification is not determined by the chemical species in the precipitation reaction.
D The chemical composition of spectator ions remains unchanged in a precipitation reaction.

Question 13 (3 MARKS)

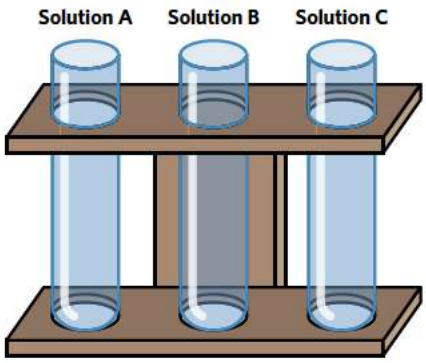
For the combination of mixtures that results in the formation of a precipitate, write out the full equation, the ionic equation and identify the spectator ions.

Exam-style questions

Within lesson

Question 14 (9 MARKS)

Some solutions that contain ionic compounds are colourless. One way of determining the identity of a solution is to use precipitation reactions. An industrial chemist has set up three test tubes, each containing an ionic solution of an unknown identity.



The chemist knows that one solution contains $Cl^-_{(aq)}$ ions, another contains $NO_3{}^-_{(aq)}$ ions and another contains $SO_4{}^{2-}_{(aq)}$ ions.

a What should the industrial chemist observe when a precipitation reaction occurs in a test tube? (1 MARK)

b Using a dropper (also known as a Pasteur pipette), the chemist places a few drops of $Ca^{2+}_{(aq)}$ into each test tube. A visual change occurs in the test tube containing solution B. Identify the anion present in solution B. Justify your answer. (3 MARKS)

c Write the ionic equation occurring in the test tube containing solution B. (1 MARK)

d Using a dropper (also known as a Pasteur pipette), the chemist places a few drops of $Ag^{+}_{(aq)}$ into each test tube. A visible change occurs in the test tube containing solution A. Identify the anions present in solution A and C and justify your answer. (3 MARKS)

e Write the ionic equation occurring in the test tube containing solution A. (1 MARK)

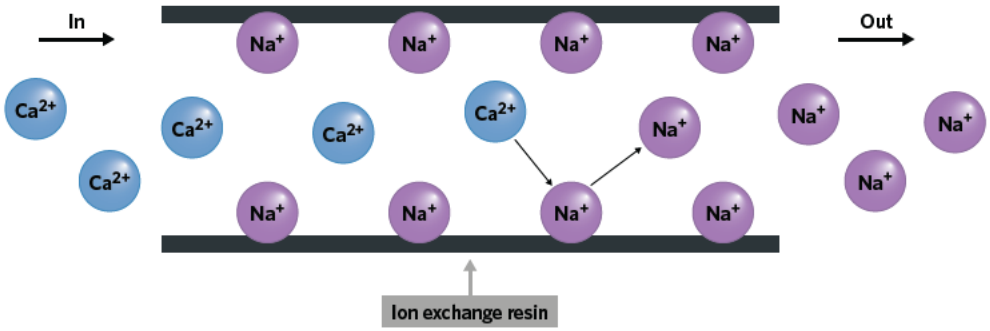
Question 15 (7 MARKS)

Limescale is a chalky and solid substance that often accumulates inside pipes and appliances like kettles. It results from a precipitation reaction that produces calcium carbonate ($CaCO_{3(s)}$).

a Write out the ionic equation for the formation of the calcium carbonate precipitate. (1 MARK)

b Pipes are often more frequently blocked with water supplies that have a high concentration of $Ca^{2+}_{(aq)}$. Suggest a reason for this observation. (2 MARKS)

One of the ways that has been invented to solve the problem of limescale is to use ion-exchange resin.



The ion exchange resin replaces the $Ca^{2+}_{(aq)}$ ions in the water with $Na^{+}_{(aq)}$ ions.

c Explain why this reduces the accumulation of limescale. (2 MARKS)

Limescale can also be removed by reacting the $CaCO_{3(s)}$ with a strong acid such as $HCl_{(aq)}$, according to the following chemical equation:

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

d Explain why the above mechanism reduces blockage in pipes caused by limescale. (2 MARKS)

Question 16 (9 MARKS)

Kidney stones are a common medical condition which occur when calcium-based precipitates are produced in the kidneys. The main precipitate is formed from calcium ions ($Ca^{2+}_{(aq)}$) and oxalate ions ($C_2O_4^{2-}{}_{(aq)}$). These precipitates can accumulate into stones and sometimes require surgery to remove.

a Write the ionic equation for the formation of kidney stones. (1 MARK)

b Dehydration increases the chance of developing kidney stones. Explain this with reference to the solubility of precipitates. (2 MARKS)

c High levels of calcium in the kidney increase the chance of developing kidney stones. Explain this with reference to the ionic equation. (2 MARKS)

A doctor specialising in kidneys (a nephrologist) wanted to determine the concentration of dissolved calcium in a patient's urine to assess the patient's risk for developing kidney stones. Urine is a mixture of many different ions including $Ca^{2+}_{(aq)}$, $Na^{+}_{(aq)}$ and $K^{+}_{(aq)}$.

d Would the presence of sodium and potassium ions affect the ability for precipitation reactions involving $Ca^{2+}_{(aq)}$ to be used? Explain your answer. (2 MARKS)

e A few drops of $Na_2CO_{3(aq)}$ are added and immediately, the urine turns cloudy. Write out the full equation for the reaction that occurs and circle the spectator ions. (2 MARKS)

Question 17 (7 MARKS)

X-rays are a very useful medical imaging technique but do not provide much information about soft tissue like the stomach or intestines. In order to investigate issues relating to these parts of the body, patients are given a 'barium meal' which consists of $BaSO_{4(s)}$ precipitate.



Image: Hong xia/Shutterstock.com

a To produce a 'barium meal', barium sulfide ($BaS_{(aq)}$) is reacted with sulfuric acid ($H_2SO_{4(aq)}$). Write out the full equation and identify the spectator ions. (2 MARKS)

b Write out the ionic equation. (1 MARK)

c Barium is highly toxic. Explain why barium sulfate, as opposed to other ionic compounds containing barium, is safe for human consumption. (2 MARKS)

d Barium sulfate solutions are often referred to by patients as a 'very thick glass of milk'. With reference to the solubility and white colour of barium sulfate, explain why this is the case. (2 MARKS)

Multiple lessons

Question 18 (7 MARKS)

Antibodies are soluble compounds used by the immune system to neutralise foreign germs or viruses dissolved in the blood, also known as antigens. When an antibody binds to an antigen, an insoluble antigen-antibody complex forms.

a Using the notation A for an antibody, B for an antigen, and C for the antigen-antibody complex, write the chemical reaction that occurs including states. (1 MARK)

b Why is this considered a precipitation reaction? (2 MARKS)

The formation of the insoluble compound is often used for rapid diagnosis of contagious diseases. A researcher at the Carlton Institute of Immunology has discovered a specific antibody for COVID-19.

c The blood samples from two patients are taken. In patient A, when the antibody for COVID-19 is supplied, a solid product forms. In patient B, nothing happens. Which patient has COVID-19? Explain your answer. (2 MARKS)

d These tests are often carried out in cool environments, rather than hot ones. With reference to solubility, explain why this is the case. (2 MARKS)

Question 19 (10 MARKS)

Wastewater is chemically treated in order to avoid potential blockages in pipes and to make it safe for human consumption. Lime ($Ca(OH)_{2(aq)}$) is used to remove $Mg^{2+}_{(aq)}$ ions and soda ash ($Na_2CO_{3(aq)}$) is used to remove $Ca^{2+}_{(aq)}$.

a Write the full and ionic equation for the removal of $Mg^{2+}_{(aq)}$. (2 MARKS)

b Write the full and ionic equation for the removal of $Ca^{2+}_{(aq)}$. (2 MARKS)

c Which precipitation reaction should be carried out first? Explain your answer with reference to the importance of removing $Ca^{2+}_{(aq)}$ and $Mg^{2+}_{(aq)}$ from the water supply to prevent pipe blockages. (2 MARKS)

d A sample of 100 mL of water is taken and 5.0 g of $Na_2CO_{3(aq)}$ is added. How many moles of $Na_2CO_{3(aq)}$ is added? (2 MARKS)

e All of the $Na_2CO_{3(aq)}$ was shown to have reacted. Given that the amount of $Na_2CO_{3(aq)}$ that reacted was equal to the amount of calcium carbonate produced, what mass of $CaCO_{3(s)}$ was formed? (2 MARKS)

Key science skills

Question 20 (8 MARKS)

The precipitates of different metals often have different colours as shown by the following table:

Cobalt	Copper	Iron (II)	Iron (III)	Manganese	Nickel	Lead
Pink	Blue	Green	Reddish brown	Pink	Green	Yellow

A mining company wanted to use the different colours of precipitates formed to determine the presence or absence of different metals.

a What is the nature of the data being collected? (1 MARK)

b Using this experimental technique, it is hard to differentiate between cobalt and manganese. Explain why this is the case and what type of error it introduces. (2 MARKS)

c The mining company wanted to accurately determine the concentration of metal cations and so they decided to approximate the concentration based on the cloudiness of the solution. Comment on the precision, reliability and accuracy of this technique. (5 MARKS)

Exam-style question hints

14a: Solids appear different to aqueous solutions. **14b:** Only some ionic compounds are insoluble and therefore precipitate out of solution. **14c:** Ionic equations don't involve spectator ions. **14d:** Only some ionic compounds are insoluble and therefore precipitate out of solution. **14e:** Ionic equations don't involve spectator ions. **15a:** Ionic equations don't involve spectator ions. **15b:** Increasing the amount of reactants increases the amount of precipitate that can form. **15c:** Some carbonate salts are soluble while others are not. **15d:** The accumulation of solids in pipes causes blockages. **16a:** Ionic equations don't involve spectator ions. **16b:** Compounds can precipitate out more easily with less water present. **16c:** Increased amount of reactants increases the amount of precipitate that can form. **16d:** Some cations are highly soluble. **16e:** The full equation involves all the chemical species present. **17a:** Spectator ions remain in the aqueous state on both sides of the chemical equation. **17b:** Ionic equations don't involve spectator ions. **17c:** The body has to absorb barium for it to be toxic and some barium compounds are soluble. **17d:** Precipitates in solution have a distinct appearance. **18a:** Precipitation reactions are identified by the states of the reactants and products. **18b:** Precipitation reactions result in the formation of a solid. **18c:** The presence of an antigen indicates the presence of a virus or germ. **18d:** Temperature affects solubility. **19a:** Ionic equations don't involve spectator ions. **19b:** Ionic equations don't involve spectator ions. **19c:** No $Ca^{2+}_{(aq)}$ or $Mg^{2+}_{(aq)}$ should be present at the end of the reactions. **19d:** The number of moles depends on the mass of the substance and its molar mass. **19e:** The mass of a compound produced depends on the mole ratio of reactants to products and the molar mass of the product. **20a:** Data can involve numbers or observations. **20b:** Some precipitates might appear very similar based on colour alone. **20c:** Judging concentration based on cloudiness is highly subjective.

EXPERIMENT

SOLUBILITY CURVE OF KNO_3

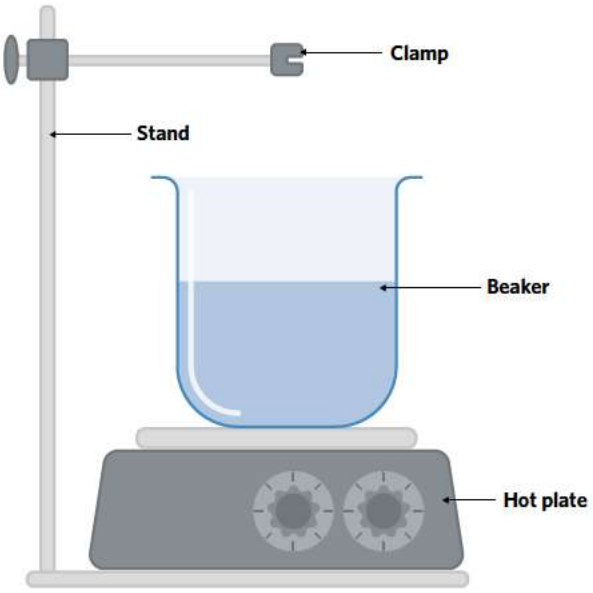
Due to their structure, ionic compounds are generally soluble in water and therefore can often be used to make different solutions. However, the amount of ionic salt (solute) that can be dissolved in water can be affected by the temperature of the solution.

In this experiment, we will be determining the solubility curve of potassium nitrate (KNO_3).

Materials

(per group)

- 1 × Hot plate
- 1 × Spatula
- 4 × Glass test tubes
- 1 × Test tube rack
- 1 × Test tube holder
- 1 × Electronic balance
- 1 × Retort stand
- 1 × Clamp
- 1 × Thermometer
- 1 × 10 mL measuring cylinder
- 1 × Stirring rod
- 400 mL tap water
- 20 mL distilled water
- 500 mL beaker
- 50.0 g potassium nitrate KNO_3 (solid)



Method

1. Label each test tube 2, 4, 6 and 8.
2. Fill the 500 mL beaker with 400 mL of tap water and place the beaker on the hot plate.
3. Set up the retort stand and clamp so that the clamp is located directly above the beaker (keep it relatively high off the beaker). This will be used to hold the test tube in place in the beaker. This setup is shown.
4. Using the hot plate, heat the water in the beaker until it reaches 90°C. Adjust the hotplate to keep the temperature constant at 90°C.
5. Measure different masses of $KNO_{3(s)}$ using the electronic balance and add a different mass to each test tube as shown in the table using the spatula.

Test tube number	Mass of $KNO_{3(s)}$ (g)
2	2.0
4	4.0
6	6.0
8	8.0

6. Using the 10 mL measuring cylinder, measure and add 5.0 mL of deionised water into each test tube.
7. Attach test tube 2 to the clamp and slowly lower into the beaker of hot water.
8. Gently stir the solution in test tube 2 using the stirring rod until the $KNO_{3(s)}$ is completely dissolved.
9. Loosen the clamp and use the test tube holder to remove the tube.
10. Add the thermometer into test tube 2 and observe the test tube until the solution begins to turn cloudy.
11. Record the temperature when the solution in test tube 2 first becomes cloudy.
12. Repeat steps 7–11 for the remaining test tubes 4, 6 and 8.

Results

Test tube number	Mass of $KNO_{3(s)}$ (g)	Temperature at which solution first becomes cloudy (°C)
2	2.0	
4	4.0	
6	6.0	
8	8.0	

QUESTIONS

Question 1 (6 MARKS)

a Convert the concentration of $KNO_{3(s)}$ in each test tube from g/5 mL to g/100 mL. (2 MARKS)

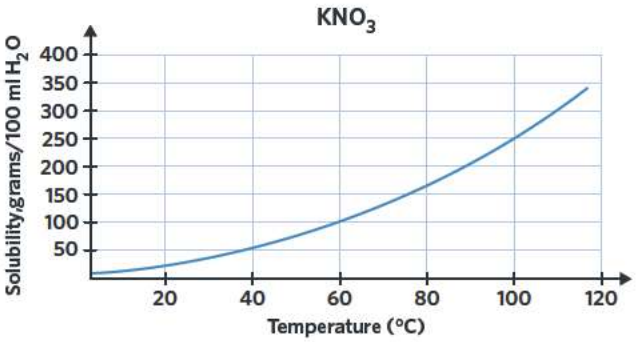
Test tube number	Mass of $KNO_{3(s)}$ (g)	Concentration of solution expressed as g/100 mL
2	2.0	
4	4.0	
6	6.0	
8	8.0	

b Using the data you have collected, draw a solubility curve for KNO_3 with the concentration of KNO_3 expressed in g/100 mL. (3 MARKS)

c Identify the trend shown by the curve. (1 MARK)

Question 2 (3 MARKS)

A student conducted the same experiment and plotted the curve shown using her results.



a Identify the solubility of KNO_3 at 90°C. (1 MARK)

b Calculate the concentration of the KNO_3 solution at 90°C, expressed in molarity (M). (2 MARKS)

ANSWERS

1 a

Test tube number	Mass of $KNO_{3(s)}$ (g)	Concentration of solution expressed as g/100 mL
2	2.0	40 g/100 mL
4	4.0	80 g/100 mL
6	6.0	120 g/100 mL
8	8.0	160 g/100 mL

b *Please check your answer with your teacher

✓ ✗ I have labeled the x-axis using the dependent variable temperature and appropriate units.

✓ ✗ I have labeled the y-axis using the independent variable concentration and appropriate units.

✓ ✗ I have plotted my curve.

c The solubility of the KNO_3 solution increases as temperature increases.

2 a 200 g/100 mL H_2O

b M is measured as mol per 1 L

$$n(KNO_3)_{100\,mL} = \frac{200}{39.1 + 14.0 + (3 \times 16.0)} = 1.978 \text{ mol}$$

$$n(KNO_3)_{1000\,mL} = \left(\frac{1000}{100}\right) \times 1.978 = 19.78 \text{ mol}$$

$$c(KNO_3) = \frac{19.78}{1} = 19.8 \text{ M}$$

CHAPTER 7 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Question 1 (1 MARK)

Which of the following characteristics is not unique to water compared to other Group 16 hydrides?

A High melting point
B High boiling point
C V-shape of the molecule
D Lower density of solid compared to liquid

Question 2 (1 MARK)

What is the energy required to completely boil 100 mL of water from room temperature at 20°C?

A 33 kJ
B 244 kJ
C 278 kJ
D 4420 kJ

Question 3 (1 MARK)

Which of the following equations does not describe water acting as a solvent?

A $AgNO_{3(s)} \rightarrow Ag^+_{(aq)} + NO_3^-{}_{(aq)}$
B $C_2H_5OH_{(l)} \xrightarrow{H_2O_{(l)}} C_2H_5OH_{(aq)}$
C $MgCl_{2(s)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)}$
D $C_5H_{11}COOC_5H_{11(l)} + H_2O_{(l)} \rightarrow C_5H_{11}COOH_{(l)} + C_5H_{11}OH_{(l)}$

Question 4 (1 MARK)

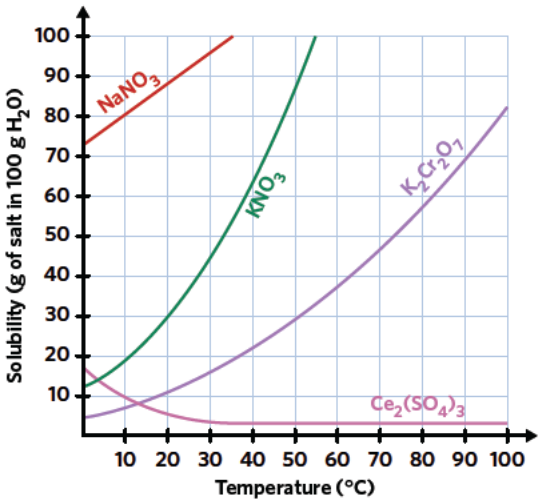
A student has 0.25 g of $Ca(OH)_{2(s)}$ and would like to generate a calcium hydroxide solution of 0.25 M. What volume of distilled water is required to prepare this solution?

A 25 mL
B 100 mL
C 13 mL
D 22 mL

Question 5 (1 MARK)

Using the solubility curve below, identify which solution is supersaturated.

A 45 g of KNO_3 in 150 g of H_2O at 30°C
B 120 g of $K_2Cr_2O_7$ in 200 g of H_2O at 90°C
C 8 g of $Ce_2(SO_4)_3$ in 100 g of H_2O at 10°C
D 50 g of $NaNO_3$ in 50 g of H_2O at 20°C



Question 6 (1 MARK)

Which of the following compounds is not soluble in water?

A Silver chloride
B CH_3OH
C KNO_3
D HCl

Question 7 (1 MARK)

Dilutions are sometimes performed because they may be too concentrated for analysis. A student diluted a sample of DNA by 1/30 (i.e. the volume of diluted DNA solution is 30 times the volume of the original solution). They then placed it in a device called a spectrophotometer, which calculated a concentration of 0.00130%(w/v).
What is the concentration of DNA in the original solution in µg/mL?

A 0.0390 µg/mL
B 0.390 µg/mL
C 3.90 µg/mL
D 390 µg/mL

Question 8 (1 MARK)

Which of the following reactions is not a precipitation reaction?

A $Pb(NO_3)_{2(aq)} + 2NaI_{(aq)} \rightarrow PbI_{2(s)} + 2NaNO_{3(aq)}$
B $2Mg_{(s)} + O_{2\,(g)} \rightarrow 2MgO_{(s)}$
C $2Na^+_{(aq)} + SO_4{}^{2-}{}_{(aq)} + Fe^{2+}{}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2(s)} + Na_2SO_{4(aq)}$
D $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$

Question 9 (1 MARK)

Which of the following statements regarding solubility is false?

A Stirring a solution increases the solubility of the solute.
B Cooling a mixture decreases the solubility of the solute.
C Changing the solvent may affect the solubility of the solute.
D A supersaturated solution is unstable.

Question 10 (1 MARK)

A research scientist was investigating the solubility of a substance at different temperatures. She knows that the following steps are involved:

I Record the temperature at which crystals start to form.

II Weigh out several samples, ensuring that each sample contains the same mass of solute and has a constant volume of water.

III Cool the test tubes slowly, stirring constantly.

IV Heat the test tubes until all solute is dissolved.

Which of the following correctly lists the order of these steps?

A III, I, II, IV

B II, III, I, IV

C II, IV, III, I

D II, III, IV, I

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (15 MARKS)

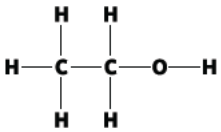
Water is often regarded as a chemical compound that is essential to life due to its vast number of different and unique properties.

a What chemical property of water gives rise to these unique properties compared to other Group 16 hydrides? Identify three unique properties and explain how they arise. (3 MARKS)

Water is important for the body's process of thermoregulation during exercise as exercising muscles produce a large amount of heat.

b An athlete's muscles have generated 100 kJ of heat energy during a period of intense physical activity. If all heat loss is through the evaporation of water from its liquid state to gaseous state, how much water (in mL) must have been perspired? (Density of water = 0.997 g/mL). (3 MARKS)

Water is also important as a biological solvent and transports many different compounds in the bloodstream, such as ethanol (pictured).



c Write the equation for the dissolution of ethanol in water. With reference to a diagram, explain why ethanol is able to dissolve in water. (3 MARKS)

d As a fully licenced driver, the legal ethanol concentration allowed in the blood is 0.05 g/100 mL. Express this value as a molar concentration. (2 MARKS)

e 98.5% of oxygen is transported by a protein called hemoglobin while 1.5% of oxygen is transported in the blood, dissolved. With reference to the polarity of oxygen, why does it require proteins to transport it around the body, as opposed to being directly dissolved in blood, most of which consists of water? (2 MARKS)

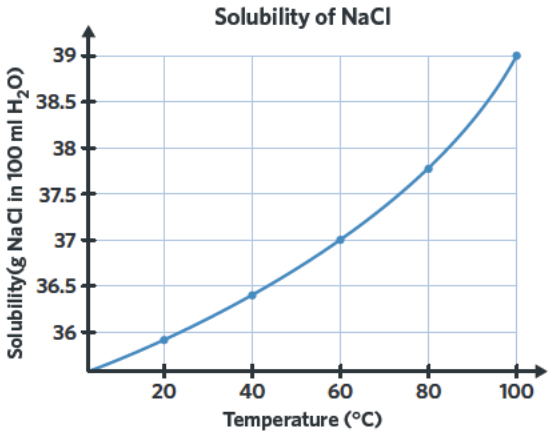
Water is also important for the preservation of aquatic life, which requires a fairly constant temperature.

f Given the oceans of the world cover many different climatic regions, explain how the thermal properties of water, as well as the density of ice, are able to maintain the temperature of water fairly constantly throughout the world. (2 MARKS)

Question 12 (15 MARKS)

Sodium chloride is a very common ionic compound that is found in table salt.

The solubility curve of sodium chloride is pictured.



a Write the equation for the dissociation of sodium chloride in water. With reference to a diagram, how does sodium chloride dissolve in water? (3 MARKS)

b What is a supersaturated solution? (1 MARK)

c Identify the concentration (in g/100 g H_2O) of NaCl required to create a supersaturated solution at 80°C and the appearance of this solution if it were disturbed. (2 MARKS)

d Identify the concentration of NaCl required to create a saturated solution at 60°C. Express this value in molarity (M). (2 MARKS)

e Based on your answer to part d, how many grams of NaCl are required to create a saturated solution with a volume of 193 mL at 60°C? (2 MARKS)

f Solutions of sodium chloride and silver nitrate are mixed together. With reference to the full chemical equation, identify the spectator ions, the precipitate formed and the balanced ionic equation for this reaction. (3 MARKS)

g Identify two experimental observations that could be made during the formation of a precipitate. (2 MARKS)

KEY SCIENCE SKILLS QUESTIONS (8 MARKS)

Question 13 (8 MARKS)

Po is investigating the effect of temperature changes on the solubility of helium gas in an aqueous solution. He would like to create a solubility curve in order to interpret the results of his experiment. Therefore, he measures the solubility of helium gas at different temperatures.

a What are the independent and dependent variables in this experiment? (2 MARKS)

b What variables should be controlled throughout the experiment? (2 MARKS)

An excerpt from the MSDS (Material Safety Data Sheet) of helium gas is shown.



Signal Words: **Warnings**

Hazard statement(s): **H280: Contains gas under pressure; may explode if heated.**

c Identify two precautions that should be taken in the laboratory to minimise the danger of working with helium gas? (2 MARKS)

d After his experiment, his results show that the solubility of helium gas in water decreases as the temperature increases. Po notes that this is the opposite to solid solutes. Suggest a plausible explanation for why the solubility of gases decreases with increasing temperature. (2 MARKS)

UNIT 2 AOS 1, CHAPTER 8

Acids and bases

08

8A An introduction to acids and bases

8B Measuring acidity

8C Reactions of acids

8D Acid-base chemistry in the body

Key knowledge

- the Brønsted-Lowry theory of acids and bases including polyprotic acids and amphiprotic species, and writing of balanced ionic equations for their reactions with water including states
- the ionic product of water, the pH scale and the use of pH in the measurement and calculations of strengths of acids and bases and dilutions of solutions (calculations involving acidity constants are not required)
- the distinction between strong and weak acids and bases, and between concentrated and dilute acids and bases, including common examples
- the reactions of acids with metals, carbonates and hydroxides including balanced full and ionic equations, with states indicated
- the causes and effects of a selected issue related to acid-base chemistry

Image: Fotaro1965/Shutterstock.com



8A AN INTRODUCTION TO ACIDS AND BASES

In this lesson, we will be learning about the definition, nature and strength of acids and bases.

8A An introduction to acids and bases	8B Measuring acidity	8C Reactions of acids	8D Acid-base chemistry in the body

Study design dot points

- the Brønsted-Lowry theory of acids and bases including polyprotic acids and amphiprotic species, and writing of balanced ionic equations for their reactions with water including states
- the distinction between strong and weak acids and bases, and between concentrated and dilute acids and bases, including common examples

Key knowledge units

The Brønsted-Lowry theory	2.1.6.1
Types of acids and bases	2.1.6.2
Strength of acids and bases	2.1.8.1

Key terms and definitions

Acid species that donates a proton, H^+, during an acid-base reaction

Base species that accepts a proton, H^+, during an acid-base reaction

Alkali base that is soluble in water

Acid-base reaction chemical reaction where a proton, H^+, is transferred between two chemical species

Ionisation process by which a chemical species gains or loses an electric charge

Dissociation process by which a molecule separates into smaller particles

Hydroxide ion negatively charged chemical species with molecular formula OH^-

Hydronium ion positively charged chemical species with molecular formula H_3O^+

Brønsted-Lowry theory definition of an acid as a chemical species which donates a proton, H^+, to a base

Conjugate acid chemical species formed from a base that has accepted a proton, H^+

Conjugate base chemical species formed from an acid that has donated a proton, H^+

Conjugate acid-base pair acid and its conjugate base or a base and its conjugate acid

Deprotonation loss of a proton, H^+, from an acid

Amphiprotic species that can either donate or accept a proton, H^+

Monoprotic species that donates one proton, H^+, during ionisation

Polyprotic species that donates multiple protons, H^+, during ionisation

Diprotic species that donates two protons, H^+, during ionisation

Triprotic species that donates three protons, H^+, during ionisation

Strength how readily an acid or base will donate or accept a proton, H^+

The Brønsted-Lowry theory 2.1.6.1

OVERVIEW

The Brønsted-Lowry theory of acids and bases defines an acid-base reaction as the donation of a proton, H^+, from an acid to a base.

THEORY DETAILS

Acids and bases are some of the most recognisable chemicals in both our everyday lives and the wider world of industrial chemistry. Since these substances are so widespread, it is important that we understand and are able to define how they function in chemical reactions. The names and molecular formulas of a number of common acids and bases are given in table 1.

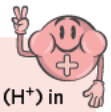
Table 1 The names and molecular formulas of common acids and bases.

Acids		Bases	
Name	Formula	Name	Formula
Sulfuric acid	H_2SO_4	Sodium hydroxide	NaOH
Hydrochloric acid	HCl	Ammonia	NH_3
Nitric acid	HNO_3	Magnesium hydroxide	$Mg(OH)_2$
Phosphoric acid	H_3PO_4	Potassium hydroxide	KOH
Ethanoic acid	CH_3COOH	Sodium carbonate	Na_2CO_3
Carbonic acid	H_2CO_3	Oxide	O^{2-}
Ammonium	NH_4^+	Calcium hydroxide	$Ca(OH)_2$

A simple definition of an acid is *'a substance that tastes sour'* and of a base, *'a substance that tastes bitter'*. This definition would seem to hold for acids such as the citric acid found in lemons or for the sodium hydroxide present in a bar of soap. However, we run into problems once we try to determine the acidity of concentrated sulfuric acid solution, H_2SO_4, or if we try and taste the basicity of ammonia, NH_3, a powerful household cleaner. In fact, it is likely that the corrosive nature of these substances would prevent your tongue from tasting anything at all.

Fortunately, we can investigate a much safer definition of acids and bases: the **Brønsted-Lowry theory**. According to this theory:

- An **acid** is a substance that donates a proton, H^+
- A **base** is a substance that accepts a proton, H^+
- An **acid-base reaction** involves the transfer of a proton from an acid to a base



To remember what happens to a proton (H^+) in an acid-base reaction, use the acronym: **B.A.A.D** **B**ases **A**ccept **A**cids **D**onate

An example of an acid-base reaction is the reaction between nitric acid, HNO_3, and water:

$$HNO_{3(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$$

$$HNO_{3(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$$

We can see that in this reaction, the nitric acid has donated a H^+ to the water molecule, forming a **hydronium ion**, H_3O^+ and a nitrate ion, NO_3^-. Using the Brønsted-Lowry theory, we can confirm that the HNO_3 acted as a proton donor, or acid, and that the H_2O acted as a proton acceptor, or base. Even though we usually think of H_2O as being a 'neutral' molecule in terms of acid-base chemistry, it is important to note that it can act as either an acid or base.

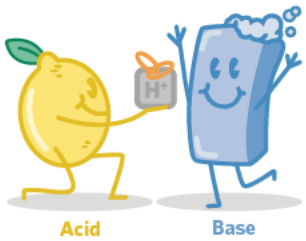


Figure 1 Acids are proton donors and bases are proton acceptors.

The **ionisation** reaction of HNO_3 in water can also be written without the presence of H_2O as shown.

$$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$$

In this scenario, the $H_3O^+_{(aq)}$ hydronium ion is instead represented as a $H^+_{(aq)}$ ion.

Throughout this chapter, these two representations are essentially identical and will only affect whether or not $H_2O_{(l)}$ is included explicitly as a reactant.

We can also use the example of the base ammonia, NH_3, reacting with water:

$$NH_{3(aq)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

$$NH_{3(aq)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

In this reaction, the ammonia has accepted a H^+ from the water molecule to form ammonium, NH_4^+ and the **hydroxide** ion, OH^-. Again, using the Brønsted-Lowry theory, since we know that NH_3 acted as a proton acceptor, it must have been a base and H_2O, since it acted as a proton donor, was the acid in this reaction.

Hydronium and hydroxide

Whenever there is a transfer of a proton, H^+, we know that an acid-base reaction must have taken place. If these reactions are happening between an acid or base and water, then either the hydronium, $H_3O^+_{(aq)}$, or hydroxide, $OH^-_{(aq)}$, ion is produced. Any base that can dissolve in water is known as an **alkali**. As shown in figure 2, the $H_3O^+_{(aq)}$ and $OH^-_{(aq)}$ ions are formed whenever $H_2O_{(l)}$ either gains or loses a proton.



The hydronium ion, $H_3O^+_{(aq)}$ and $H^+_{(aq)}$ can be used interchangeably for acid-base reactions.



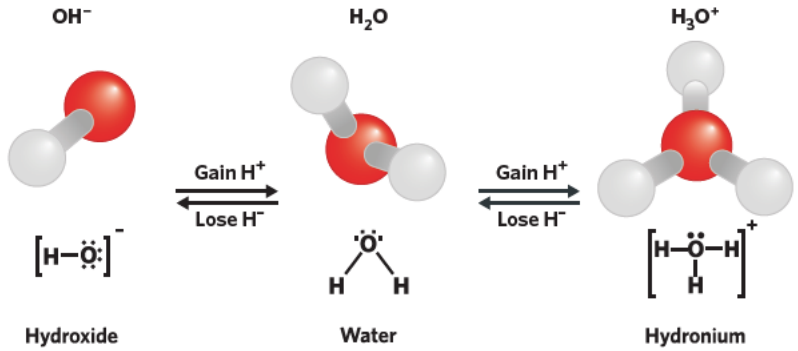


Figure 2 The hydroxide and hydronium ions are formed from the transfer of a proton and a water molecule.

Not all acid-base reactions involve water as a reactant. In fact, we could just as easily have a reaction between $HNO_{3(aq)}$ and $NH_{3(aq)}$. In this reaction nitric acid will act as an acid and ammonia as a base through the transfer of a proton:

$$HNO_{3(aq)} + NH_{3(aq)} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$$

$$HNO_{3(aq)} + NH_{3(aq)} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$$

Worked example 1

Write the balanced equation for the acid-base reaction of hydrochloric acid, HCl, in aqueous solution.

What information is presented in the question?

The molecular formula of hydrochloric acid, HCl.

The reaction is occurring in aqueous solution.

What is the question asking us to do?

Write a balanced equation for the acidic behaviour of HCl in water.

What strategy(ies) do we need in order to answer the question?

1. Remember that aqueous solution means that the reaction is occurring in water.
2. Understand that an acid will donate a H^+ ion during an acid-base reaction.

Answer

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + Cl^-{}_{(aq)}$$

Acid Base

As an acid, $HCl_{(aq)}$ will donate a H^+ ion to the base in an acid-base reaction. Since this reaction is occurring in aqueous solution, $H_2O_{(l)}$ is present and can act as a base. The $H_2O_{(l)}$ will accept the H^+ ion from HCl, resulting in the formation of the hydronium ion, $H_3O^+{}_{(aq)}$. The $HCl_{(aq)}$, since it ionises to donate its H^+ ion, will become an aqueous chloride ion, $Cl^-{}_{(aq)}$.

This gives the full overall balanced equation as:

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + Cl^-{}_{(aq)}$$

Conjugate acid-base pairs

When an acid donates a H^+ ion to a base, we can imagine that a 'space' for a H^+ ion is created in the acid molecule when it leaves. It is possible for that H^+ ion to combine again with the acid molecule. If it does, then the acid molecule would be accepting a H^+ ion. However, this is our Brønsted-Lowry theory definition of a base! Since the molecule that remains after an acid donates a H^+ ion is a base, we define this molecule as the **conjugate base** of that particular acid. An acid and its conjugate base are, together, defined as a **conjugate acid-base pair**.

This is also true for when a base accepts a H^+ ion during an acid-base reaction. Suddenly, it has one extra H^+ ion that it can donate, meaning it can function as an acid. This molecule formed from a base that has accepted a H^+ ion is, therefore, defined as the **conjugate acid** of the base. Together, these two species also constitute a conjugate acid-base pair.

The reactions of HNO_3 and NH_3 in water and with each other are given below, with all conjugate acid-base pairs shown in figure 3.



Conjugate acid-base pairs can easily be recognised because their molecular formulas will differ by one H^+. Remember that the acid or conjugate acid will always be the species that has one more H^+.

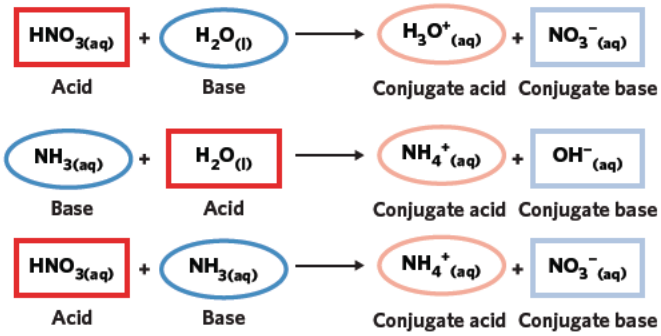


Figure 3 Acid-base reactions and their conjugate acid-base pairs.

Worked example 2

The equation for the acid-base reaction between ethanoic acid, CH_3COOH, and ammonia, NH_3, is shown.

$CH_3COOH_{(aq)} + NH_{3(aq)} \rightarrow CH_3COO^-_{(aq)} + NH_4^+{}_{(aq)}$

Identify the conjugate acid-base pairs in this reaction.

What information is presented in the question?

The balanced acid-base reaction between CH_3COOH and NH_3.

What is the question asking us to do?

Identify the conjugate acid-base pairs present in this reaction.

What strategy(ies) do we need in order to answer the question?

1. Identify which species is the acid and which species is the base.
2. Determine the conjugate base and acid for these two species by looking for ions that would form from either an accepted or donated H^+ ion.

Answer

$\underset{\text{Acid}}{CH_3COOH_{(aq)}} + \underset{\text{Base}}{NH_{3(aq)}} \rightarrow \underset{\text{Conjugate base}}{CH_3COO^-_{(aq)}} + \underset{\text{Conjugate acid}}{NH_4^+{}_{(aq)}}$

Looking at the CH_3COOH on the left hand side of the equation and the CH_3COO^- on the right, we know that CH_3COOH must have donated a H^+ ion. As a result, CH_3COOH must be the acid and CH_3COO^- is its conjugate base as they differ only by a H^+ ion.

Similarly, the NH_3 on the left has accepted a H^+ ion to form NH_4^+, acting as a base and making NH_4^+ its conjugate acid.

Therefore, the conjugate acid-base pairs are:

$CH_3COOH_{(aq)}$ / $CH_3COO^-_{(aq)}$

$NH_4^+{}_{(aq)}$ / $NH_{3(aq)}$

Types of acids and bases 2.1.6.2

OVERVIEW

Acids and bases can be classified according to their capacity to donate or accept a number of protons.

THEORY DETAILS

Amphiprotic species

We have established that for a chemical species to be classified as an acid, it must be able to donate a proton, H^+, and to be a base, it must be able to accept a proton. However, in the previous section of this lesson, we investigated the following two reactions.

$\underset{\text{Acid}}{HNO_{3(aq)}} + \underset{\text{Base}}{H_2O_{(l)}} \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$

$\underset{\text{Base}}{NH_{3(aq)}} + \underset{\text{Acid}}{H_2O_{(l)}} \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$

In these two acid-base reactions, we can see that $H_2O_{(l)}$ can be acting either as a base or as an acid depending on which species it is reacting with. Any substance that can act as either an acid or a base is defined as an **amphiprotic** substance, meaning that it can both donate and accept a proton in an acid-base reaction.

A list of amphiprotic substances and the ions they form when acting as an acid or a base is given in figure 4. All amphiprotic substances have both a 'space' to accept an H^+ ion, as well as an H^+ ion they are prepared to donate.

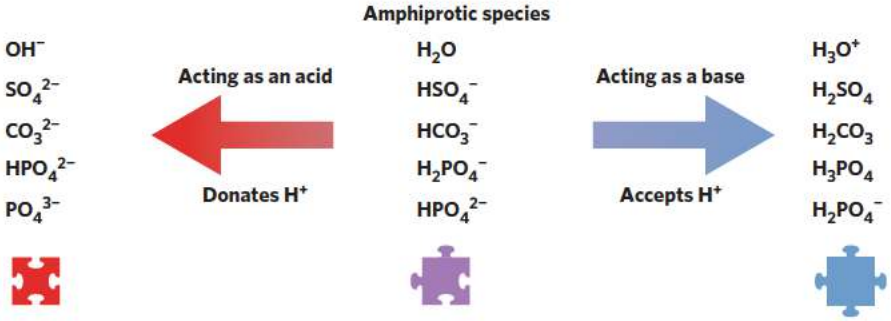


Figure 4 Amphiprotic species can donate or accept a H^+, acting as either an acid or a base.

The classification of acids

We know that any species that can donate a H^+ is classified as an acid. We can define this process of giving up a proton as **deprotonation**. Any acid that will only donate one proton is classified as **monoprotic**. Once these acids donate their only H^+, they can no longer react as acids in acid-base reactions. We have already investigated the ionisation reaction of nitric acid, HNO_3, and, removing $H_2O_{(l)}$ from the equation, we can more closely follow nitric acid's behaviour.

$$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$$

Once $HNO_{3(aq)}$ gives up its only H^+, the nitrate ion that forms has no more H^+ ions to donate. Therefore, we would classify HNO_3 as a monoprotic acid.

There are a number of acids that are capable of undergoing deprotonation multiple times. This means that they can donate a H^+ and then the compound formed from this deprotonation can again donate a H^+. These species are defined as being **polyprotic**, meaning they can donate a H^+ more than once. If an acid can donate two H^+ ions, then it is classified as **diprotic**. An example of a diprotic acid is sulfuric acid, H_2SO_4. Sulfuric acid ionises in water according to the following equations.

1 $H_2SO_{4(aq)} \rightarrow HSO_4^-{}_{(aq)} + H^+_{(aq)}$

2 $HSO_4^-{}_{(aq)} \rightleftharpoons SO_4^{2-}{}_{(aq)} + H^+_{(aq)}$

First, one proton is lost to form HSO_4^-, and then another is donated to form the sulfate ion, SO_4^{2-}.

Note that when H_2SO_4 ionises in water, this second deprotonation step only occurs to about 10% of the extent of the first stage. However, when H_2SO_4 reacts with a base, such as NH_3, this second proton is lost to a much greater (almost complete) extent.

If an acid can donate three H^+ ions, then it is classified as **triprotic**. A common example of a triprotic acid is phosphoric acid, H_3PO_4. Phosphoric acid deprotonates in three stages, since it has the potential to lose three H^+ ions. These reactions are given below.

1 $H_3PO_{4(aq)} \rightarrow H_2PO_4^-{}_{(aq)} + H^+_{(aq)}$

2 $H_2PO_4^-{}_{(aq)} \rightleftharpoons HPO_4^{2-}{}_{(aq)} + H^+_{(aq)}$

3 $HPO_4^{2-}{}_{(aq)} \rightleftharpoons PO_4^{3-}{}_{(aq)} + H^+_{(aq)}$

However, since phosphoric acid is relatively weak, the second and third deprotonation steps here only occur to a low extent. We can use 'bidirectional harpoon' arrows (⇌) to show that this reaction will have a large proportion of reactants that haven't ionised. A simple way to explain this is that once a polyprotic acid loses an H^+ ion, it becomes negatively charged. Since opposite charges attract, it then becomes much harder for the positive H^+ ion to escape the pull of the full compound's negative charge. As a result, the second and third deprotonation occur to an increasingly lower extent than the first.

If we were to have a look at a solution of phosphoric acid in water, rather than seeing only PO_4^{3-} and H^+ ions, we would instead see a mixture of H_3PO_4, $H_2PO_4^-$, HPO_4^{2-}, PO_4^{3-} and H^+ ions, as shown in figure 5.

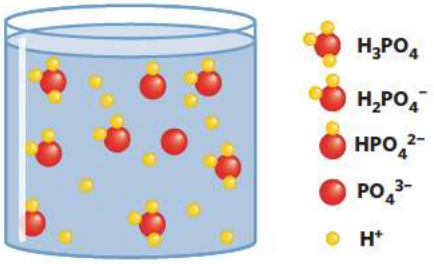


Figure 5 A solution of phosphoric acid will contain a mixture of many different ions.

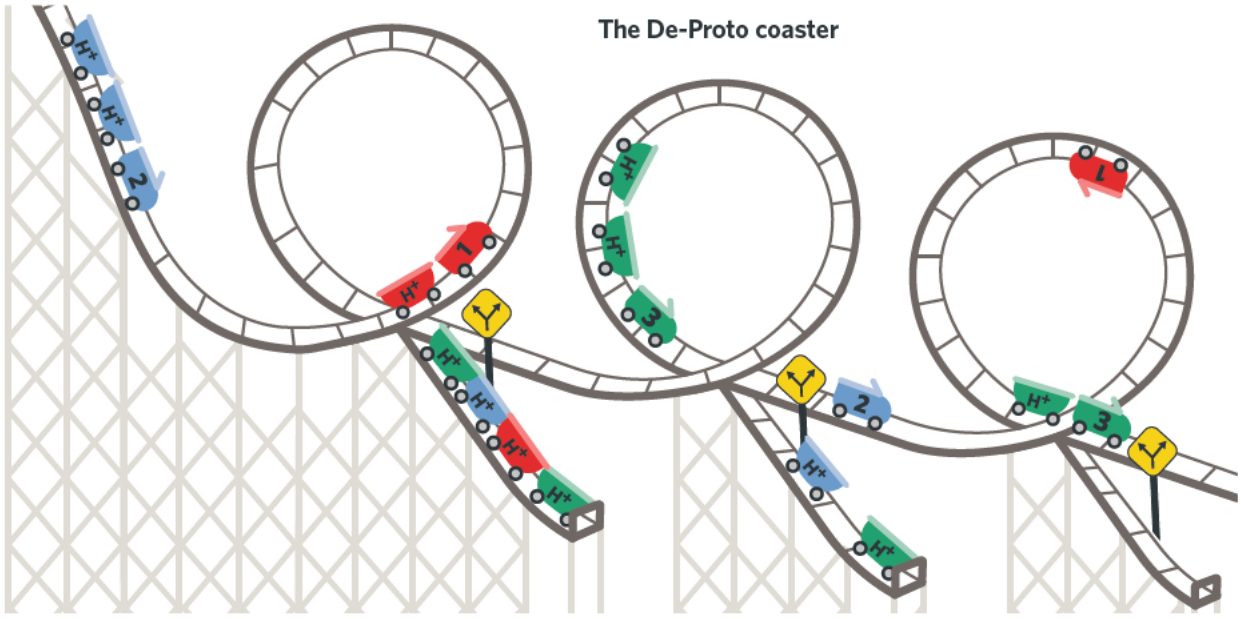


Figure 6 Polyprotic acids donating their H^+ ions can be thought of as a roller coaster and a number of loop-de-loops. After each loop-de-loop, an H^+ cart is lost, but the momentum for the carts to pass through the next loop-de-loop decreases. This is similar to how, with each deprotonation, it becomes harder for a polyprotic acid to donate another H^+ ion.

Table 2 contains a list of common acids and whether they are classified as monoprotic, diprotic or triprotic. Note that a number of amphiprotic species are formed from the deprotonation of a polyprotic acid, since any proton that is lost can be re-accepted if the polyprotic acid acts as a base in another reaction.

Table 2 List of some monoprotic, diprotic and triprotic acids.

Monoprotic		**Diprotic**		**Triprotic**	
Name	Formula	Name	Formula	Name	Formula
Nitric acid	HNO_3	Sulfuric acid	H_2SO_4	Phosphoric acid	H_3PO_4
Hydrochloric acid	HCl	Carbonic acid	H_2CO_3	Citric acid	$C_6H_8O_7$
Ethanoic acid	CH_3COOH	Hydrogen sulfide	H_2S	Boric acid	H_3BO_3
Ammonium	NH_4^+	Selenous acid	H_2SeO_3	Arsenic acid	H_3AsO_4
Hydrofluoric acid	HF				

Strength of acids and bases 2.1.8.1

OVERVIEW

The strength of an acid or base is related to how willingly they donate or accept a H^+ ion.

THEORY DETAILS

So far, we have classified acids and bases according to whether they donate or accept H^+ ions and how many H^+ ions they are capable of donating. We will now investigate the **strength** of acids and bases, which depends on how readily an acid donates a H^+ ion or how readily a base accepts a H^+ ion.

Acids

A strong acid is one that will undergo almost complete ionisation when added to water. This means that essentially all of the acid molecules in the solution will donate their protons to form aqueous H^+ ions. There are three common strong acids, all of which will lose their first proton very readily in aqueous solution.

Sulfuric acid: $H_2SO_{4(aq)} \rightarrow HSO_4^-{}_{(aq)} + H^+{}_{(aq)}$

Hydrochloric acid: $HCl_{(aq)} \rightarrow Cl^-{}_{(aq)} + H^+{}_{(aq)}$

Nitric acid: $HNO_{3(aq)} \rightarrow NO_3^-{}_{(aq)} + H^+{}_{(aq)}$

For example, in a 1.00 M solution of HCl, well over 99.9% of HCl molecules will have ionised completely into $Cl^-_{(aq)}$ and $H^+_{(aq)}$ ions. For most purposes, zero molecules of un-ionised HCl remain in solution and so we can use a complete, unidirectional arrow (→) to show that the reaction goes to completion. The complete ionisation of HCl is depicted in figure 7.

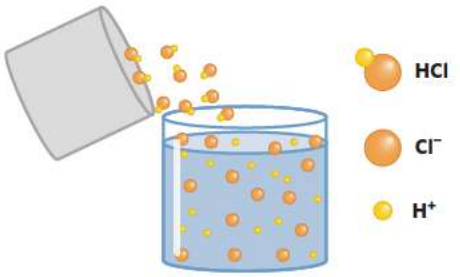


Figure 7 Solutions of $HCl_{(aq)}$ will be almost completely composed of just $H^+_{(aq)}$ and $Cl^-_{(aq)}$ ions.

A weak acid, on the other hand, is one that does not readily give up its H^+ ion in solution. The ionisation equations for three common weak acids are given.

Ethanoic acid: $CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$

Carbonic acid: $H_2CO_{3(aq)} \rightleftharpoons HCO_3^-{}_{(aq)} + H^+_{(aq)}$

Ammonium: $NH_4^+{}_{(aq)} \rightleftharpoons NH_{3(aq)} + H^+_{(aq)}$

None of these acids will ionise in water to a significant extent and so they are classified as weak acids. We can use the bidirectional harpoon arrows (⇌) to show that these reactions do not go to completion. For example, only about 0.5% of the CH_3COOH molecules in a 1.0 M solution will ionise into CH_3COO^- and H^+ ions (1 out of every 200). The rest will simply remain as unionised aqueous CH_3COOH molecules. The incomplete ionisation of $CH_3COOH_{(aq)}$ is depicted in figure 8.

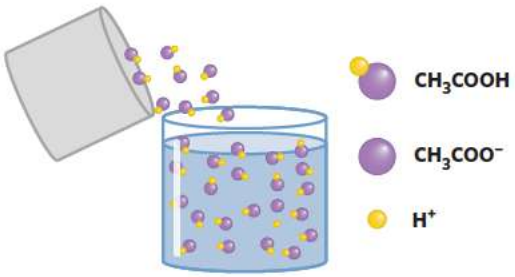
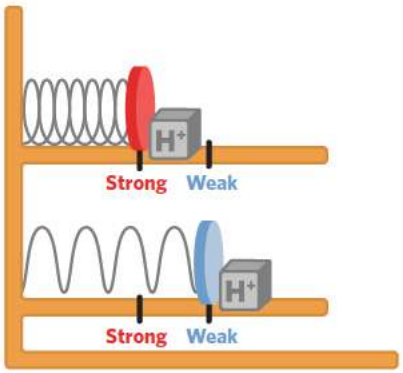


Figure 8 Ethanoic acid in aqueous solution will contain mostly $CH_3COOH_{(aq)}$ molecules, with very few $CH_3COO^-_{(aq)}$ and $H^+_{(aq)}$ ions.

As shown in figure 9, strong acids are like tightly coiled springs that will readily release a H^+ ion for an acid-base reaction. Weak acids, on the other hand, are hesitant and do not give up their protons as willingly, making them like much less compressed coiled springs.



Strong acids
Sulfuric acid H_2SO_4
Hydrochloric acid HCl
Nitric acid HNO_3

Weak acids
Ethanoic acid CH_3COOH
Carbonic acid H_2CO_3
Ammonium acid NH_4^+

Figure 9 Strong and weak acids can be compared to coiled springs compressed by different amounts.

To help remember which acids are strong and which ones are weak, use the acronym SHN | ECA.
Sulfuric Hydrochloric Nitric | Ethanoic Carbonic Ammonium
(Strong) (Weak)



Bases

The rules for classifying strong and weak bases are very similar to those used for acids, except that a base's strength depends on how readily it will accept a H^+ ion from an acid.

Examples of strong bases include the oxide ion, O^{2-}, the hydroxide ion, OH^-, and hydrogen phosphate, HPO_4^{2-}. All of these species will readily accept a H^+ ion during an acid-base reaction. The reaction of O^{2-} in water is given figure 10. This reaction will reach completion with a majority of O^{2-} ions having accepted a H^+ ion to form a OH^- ion. A second OH^- ion is formed once $H_2O_{(l)}$ donates its proton.

$O^{2-}_{(aq)} + H_2O_{(l)} \rightarrow OH^-_{(aq)} + OH^-_{(aq)}$

Since the hydroxide ion, OH^-, is such a strong base, any compound that more or less completely **dissociates** (separates into smaller particles) to form OH^- ions will create a basic solution. Ionic compounds such as NaOH, KOH and $Ca(OH)_2$ all function as strong bases since they readily dissociate to form OH^- ions. However, according to Brønsted-Lowry theory, these ionic compounds are not technically bases themselves, since they do not accept any H^+ ions directly. They instead act as a source of the strong base OH^-.

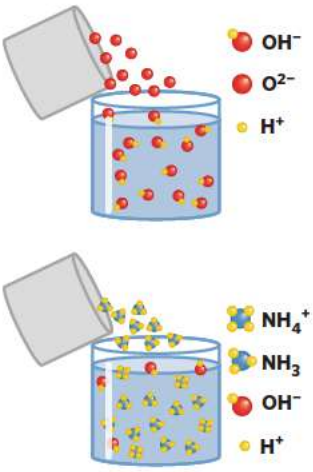


Figure 10 The distribution of ions that would be present in solutions of the strong base $O^{2-}_{(aq)}$ and the weak base $NH_{3(aq)}$.

Weak bases, on the other hand, do not readily accept H^+ ions. An example of the reaction between the weak base ammonia, NH_3, and water is given below. Again, we use bidirectional harpoon arrows to show that this reaction is only partially complete.

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

We classify ammonia as a weak base as only a small proportion of the NH_3 molecules present in the aqueous solution will accept a H^+ ion to form NH_4^+. This means that only a small fraction of the molecules present in the solution will be hydroxide ions, OH^-, which is an indicator of how basic the resulting solution is. Most of the molecules in solution will remain as unionised ammonia, NH_3.

Strong bases can be thought of as powerful magnets that readily attract and accept H^+ ions, whereas weak bases are much less willing to accept a H^+ ion and can, therefore, be compared to weaker magnets. This analogy is shown in figure 11.

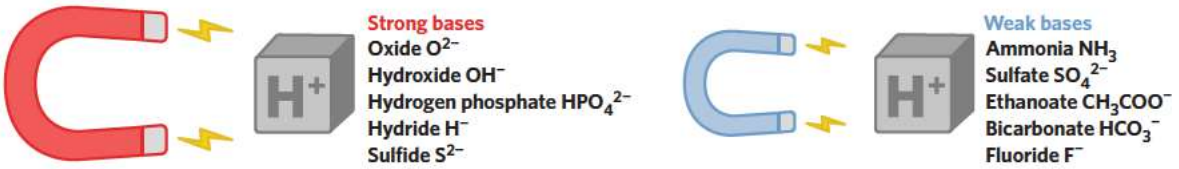


Figure 11 Strong and weak bases can be compared to magnets of different strength. The strong bases succeed much more than the weak bases in 'attracting' and accepting a H^+ ion.

Adjectives such as 'concentrated', 'dilute', 'strong' and 'weak' are all *qualitative* descriptions of an acid. Precise measurements of concentration with units such as mol L^{-1} (M) or g L^{-1} would be *quantitative* descriptions of concentration. In VCE chemistry, quantitative measurements of an acid's strength are not required.

Strength vs concentration

It is important to note the difference between the concentration of an acid or base in solution and its strength. The strength of an acid relates to what proportion of its molecules will donate a H^+ ion. Its concentration depends on how many of those acid molecules there were to begin with.

In the case of hydrochloric acid, if we use a small amount of HCl, we will end up with a dilute solution. If we use a large amount, then we will have a concentrated solution. However, both of these solutions will be solutions of the strong acid HCl.

In the case of ethanoic acid, if we use a small amount of CH_3COOH, we will still end up with a dilute solution. If we use a large amount, then we will again still have a concentrated solution. The only difference is that in both these cases, we will end up with only a small percentage of $CH_3COO^-{}_{(aq)}$ and $H^+{}_{(aq)}$ ions. However, both of these solutions will be solutions of the weak acid CH_3COOH.

Adding or removing acid molecules from solution will only change its concentration, not its strength, since an acid or base's strength depends only on the identity of the acid or base itself. Figure 12 shows common acids and solutions that would be classified as some combination of strong or weak and concentrated or dilute.

Don't confuse strength with concentration. Strength depends on what kind of acid is present. Concentration depends on how much of that acid is present. It is possible to have a dilute solution of a strong acid or a concentrated solution of a weak acid.

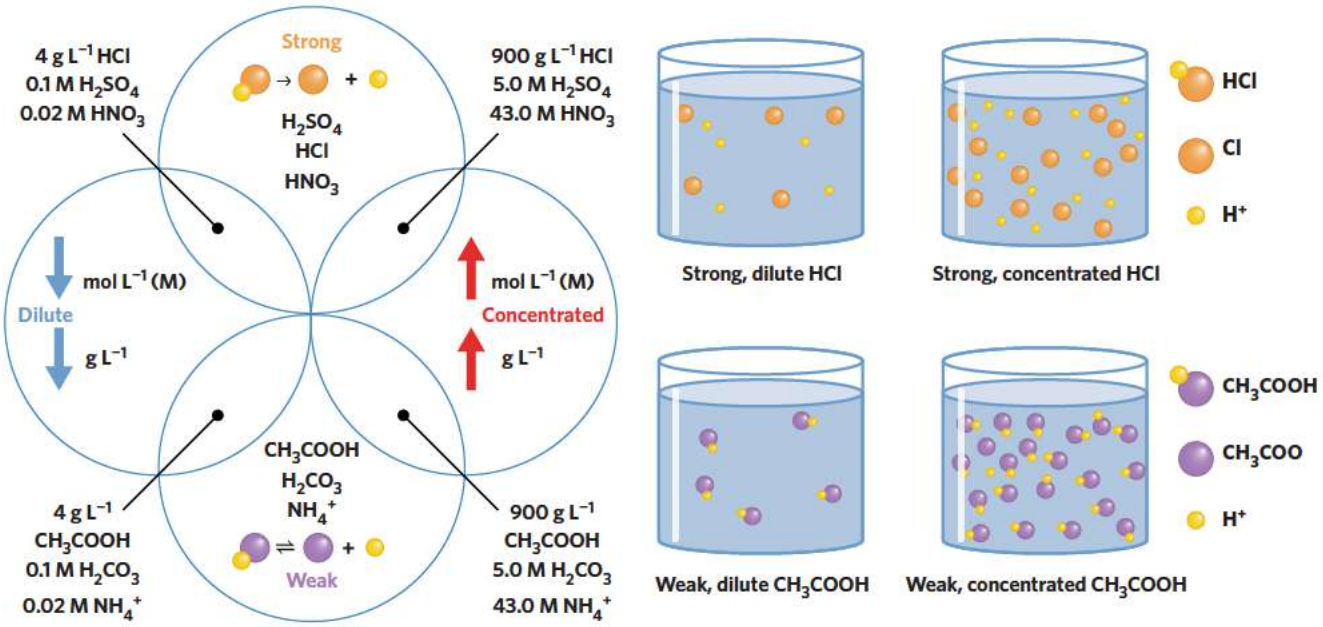


Figure 12 Venn diagram and figures showing the differences between strong, weak, concentrated and dilute acids.

Strength of conjugate acid-base pairs

When we investigated conjugate acid-base pairs, we discovered that they can be identified as species that differ by one H^+ ion. For example, consider the reaction between the weak acid hydrogen sulfide and the strong base of the hydroxide ion.

$$\underset{\text{weak acid}}{H_2S_{(aq)}} + \underset{\text{strong base}}{OH^-_{(aq)}} \rightarrow \underset{\text{conjugate acid}}{H_2O_{(aq)}} + \underset{\text{conjugate base}}{HS^-_{(aq)}}$$

The conjugate acid-base pairs in this reactions are:

- $H_2S_{(aq)} / HS^-_{(aq)}$
- $H_2O_{(l)} / OH^-_{(aq)}$

Strong acids have weak conjugate bases. This would make sense as, if a strong acid readily donates a H^+ ion, the conjugate base that forms is not going to easily re-accept that H^+ ion, making it a weak base. Likewise, strong bases have weak conjugate acids. Figure 13 shows this relationship between conjugate acid-base pairs.

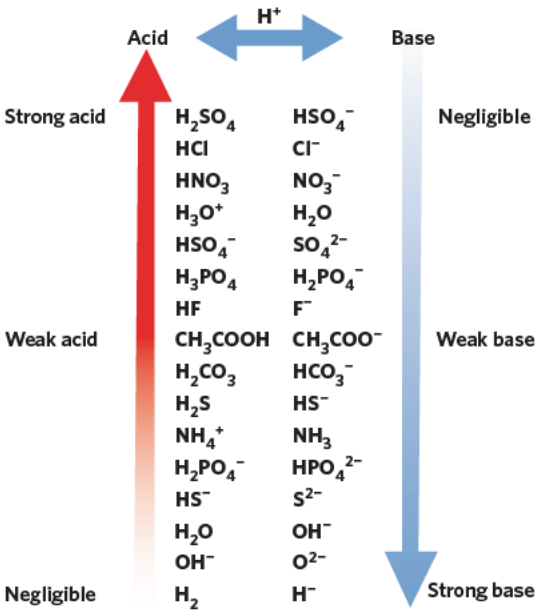


Figure 13 Relative strength of common conjugate acid-base pairs.

Theory summary

- According to the Brønsted-Lowry theory of acids and bases:
  - Acids donate H^+ ions.
  - Bases accept H^+ ions.
- The hydronium ion, $H_3O^+_{(aq)}$, and $H^+_{(aq)}$ can be used interchangeably for acid-base reactions to represent an aqueous proton.
- Conjugate acid-base pairs:
  - After an acid donates a H^+ ion, the chemical species it forms is its conjugate base.
  - After a base accepts a H^+ ion, the chemical species it forms is its conjugate acid.
  - These pairs are known as conjugate acid-base pairs.
  - The stronger the acid, the weaker its conjugate base. The stronger the base, the weaker its conjugate acid.
- Classification of acids and bases:
  - Amphiprotic species are capable of accepting or donating a H^+ ion.
  - Monoprotic acids can donate 1 proton.
  - Polyprotic acids can donate multiple protons.
  - Diprotic acids can donate 2 protons.
  - Triprotic acids can donate 3 protons.
- Each subsequent deprotonation step becomes harder and harder for polyprotic acids.
- Strong acids and bases ionise completely in water.
- Weak acids and bases ionise only partially in water.



The stronger the acid, the weaker its conjugate base. The stronger the base, the weaker its conjugate acid.

8A QUESTIONS

Theory review questions

Question 1

Acid-base reactions involve the

A transfer of neutrons.

B transfer of H^+ ions.

Question 2

Bases _____ H^+ ions, acids _____ H^+ ions.

A accept, donate

B donate, accept

Question 3



In this reaction,

A $HCl_{(aq)}$ is acting as a base, $H_2O_{(l)}$ is acting as an acid.

B $HCl_{(aq)}$ is acting as an acid, $H_2O_{(l)}$ is acting as a base.

Question 4

Hydrofluoric acid, HF, is a/an

A amphiprotic acid.

B monoprotic acid.

Question 5

Dihydrogen phosphate, $H_2PO_4^-$ is

A diprotic only.

B diprotic and amphiprotic.

Question 6

All triprotic acids are polyprotic.

A True

B False

Question 7

$H_2CO_{3(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3^-{}_{(aq)} + H_3O^+{}_{(aq)}$

In this reaction, $H_2CO_{3(aq)}$ / $HCO_3^-{}_{(aq)}$ is an example of a

A conjugate acid-base pair.

B complimentary acid-base pair.

Question 8

Which of the following is a weak acid?

A Nitric acid, HNO_3

B Ammonium, NH_4^+

Question 9

As a strong acid, a 1.0 M solution of HCl will contain

A almost zero HCl molecules.

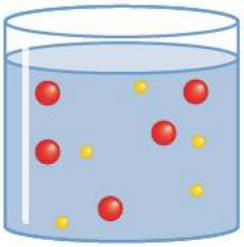
B almost zero H^+ and Cl^- ions.

Question 10

The given diagram most likely depicts a

A weak, dilute acid.

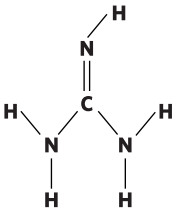
B strong, dilute acid.



Deconstructed questions

Use the following information to answer questions 11–13.

The structural formula of Guanidine, $HNC(NH_2)_2$ is given. It is a strong base often used in the manufacturing of plastic and explosives. Guanidine is capable of accepting one H^+ ion during acid-base reactions.



Question 11 (1 MARK)

Which of the following gives the molecular formula of the conjugate acid of guanidine?

A $C(NH_2)_3^+$
B H_3O^+
C $NC(NH_2)_2^-$
D H_2O

Question 12 (1 MARK)

Consider a solution of $HNC(NH_2)_{2(aq)}$. Which of the following species would most likely be present in high concentrations in this solution?

I OH^- **II** H_3O^+ **III** $NC(NH_2)_2^-$ **IV** $C(NH_2)_3^+$ **V** $HNC(NH_2)_2$

A I, III, V
B I, IV
C II, III, V
D I, IV, V

Question 13 (5 MARKS)

Guanidine is highly soluble in water because of its multiple highly electronegative nitrogen atoms.

a Give the balanced equation for the acid-base reaction between guanidine and water. (1 MARK)

b Which chemical species in this reaction is acting as the acid? Justify your answer with reference to the Brønsted-Lowry theory of acids and bases. (2 MARKS)

c Would this chemical species be classified as monoprotic, polyprotic or amphiprotic? Explain. (2 MARKS)

Exam-style questions

Within lesson

Question 14

0.03 M H_2SO_4 is an example of a

A weak, dilute acid.
B strong, concentrated acid.
C weak, concentrated acid.
D strong, dilute acid.

Use the following information to answer questions 15–16.

"Perhaps it should be mentioned that H^+ cations are unique in having no electrons. Their small size makes their charge density very high. The H^+ ion is a proton, and free protons do not really exist in water solution. Although often written as H^+ for the sake of simplicity, the hydrogen ions in acid solutions are actually attached to molecules of the solvent. One way to indicate this in equations is by writing $H^+_{(aq)}$ to signify the "aqueous" solvation of the ion. Another way is by using the hydronium ion, H_3O^+, which is a proton attached to a molecule of water.



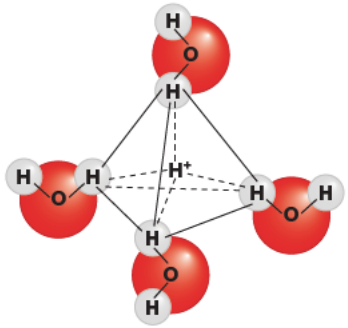
Since liquid water molecules exist in hydrogen bonded clusters, with about four H_2O units in an average group at room temperature, it may be that the actual formula for the aquated hydrogen ion is closer to $H_9O_4^+$ (although it is rarely written this way). In any case, that fact does not lessen the usefulness of the hydronium ion, which remains the simplest and most practical way to indicate the covalent attachment of H^+ ions to water."

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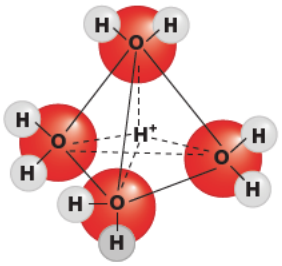
Question 15 (1 MARK)

Using the information presented in the article and your knowledge of hydrogen bonding and electronegativities, which of the following best represents the true structure of a solvated proton in aqueous solution?

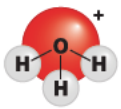
A



B



C



D



Question 16 (1 MARK)

Which of the representations of an aquated proton, $H^+_{(aq)}$, does the article suggest is the most suitable?

A H^+ because it is the simplest and most efficient way of showing protons in solution.

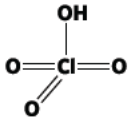
B $H_9O_4^+$ because it represents the true and most accurate structure of an aquated proton.

C H_3O^+ because it offers a structure that represents the dispersion of the proton's charge.

D H_3O^+ because it conveniently displays an aquated proton's covalent bond with the solvent molecules.

Question 17 (6 MARKS)

Perchloric acid, $HClO_4$, is a considerably stronger acid than both sulfuric acid and nitric acid. For these reasons, it is actually classified as a 'super acid'. In high concentrations, it is dangerously corrosive and is used in the synthesis of rocket fuel. Its structure is shown.



a Is perchloric acid a monoprotic, polyprotic or amphiprotic acid? (1 MARK)

b Give the balanced equation for the acid-base reaction between $HClO_4$ and the alkali NH_3 in aqueous solution. (1 MARK)

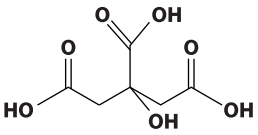
c Perchloric acid and its conjugate base together make a conjugate acid-base pair.

i Give the molecular formula for the conjugate base of perchloric acid, $HClO_4$. (1 MARK)

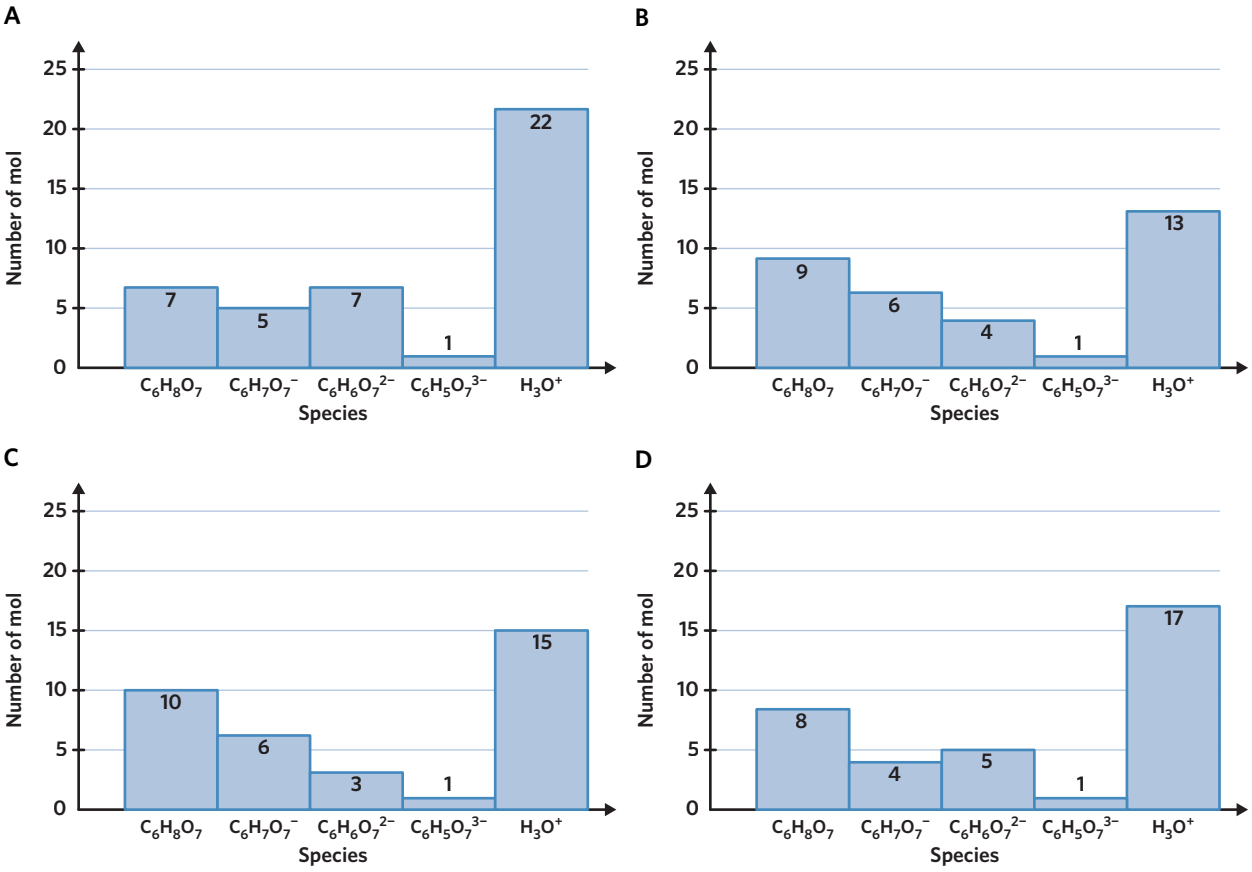
ii If a solution of this conjugate base were added to a vessel of pure water, would the concentration of $OH^-_{(aq)}$ ions increase by any noticeable extent? Justify your answer with the aid of a balanced acid-base equation. (3 MARKS)

Question 18 (2 MARKS)

Citric acid, $C_6H_8O_7$ is a relatively weak triprotic acid that is found naturally in citrus fruits. The structure of citric acid is given. 20 mol of citric acid is added to a vessel of water and left for a few minutes until no further activity is observed.



a Which of the following charts would represent the most likely distribution of aqueous species present in the resulting solution? (1 MARK)



b Which of the following best explains the distribution of the ions present in the citric acid solution? (1 MARK)

A Polyprotic acids deprotonate randomly and spontaneously, leading to a random distribution of ions.

B With each deprotonation, it becomes much harder for the positive H^+ ion to escape the pull of the remaining compound's negative charge, which leads to a progressively decreasing distribution of ions. As each proton is lost to form a H_3O^+ ion, these positive charges electrostatically resist and repel any new positive H^+ ions from entering the solution, leading to a decreasing distribution of fully deprotonated species.

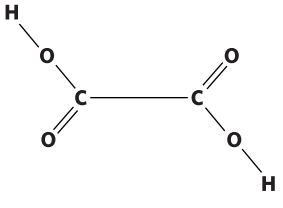
C Deprotonation results in the destabilisation of polyprot ic acids and so they resist deprotonation to maintain this stability, leading to the species in highest concentration being the complete molecule.

D Deprotonation results in the destabilisation of polyprotic acids and so they resist deprotonation to maintain this stability, leading to the species in highest concentration being the complete molecule.

Multiple lessons

Question 19 (11 MARKS)

Oxalic acid is the simplest possible dicarboxylic acid, which means that it contains two carboxyl functional groups. It has the semi-structural formula HOOCCOOH or $(COOH)_2$. Oxalic acid is a naturally occurring compound found in the flowers of the wood-sorrel or *Oxalis* genus. The structural formula of oxalic acid is shown.



a Sketch a diagram of the hydrogen bonding interaction that would occur between two molecules of oxalic acid. (2 MARKS)

b Consider a 5.0 M solution of $(COOH)_{2(aq)}$. The percentage composition by mass of the solution is shown in the given table.

Species	% composition by mass
$(COOH)_2$	25.7
$(COOH)(COO)^-$	4.30
$(COO)_2^{2-}$	0.912
H_3O^+	1.29

i Why does the '% composition by mass' column not sum to a total of 100%? (1 MARK)

ii Would oxalic best be classified as a strong or weak acid? Justify your answer with reference to the data presented in the table. (3 MARKS)

c Oxalic acid will react with the strong base $OH^-_{(aq)}$ in aqueous solution according to the following balanced reaction.

$(COOH)_{2(aq)} + 2OH^-_{(aq)} \rightarrow (COO)_2^{2-}{}_{(aq)} + 2H_2O_{(l)}$

If 5.00 L of 2.30 M oxalic acid reacts completely with $OH^-_{(aq)}$ ions, calculate

i the number of molecules of $(COOH)_{2(aq)}$ that are reacting. (1 MARK)

ii the mass, in grams, of $OH^-_{(aq)}$ ions required for this reaction. (1 MARK)

iii what kind of acid would oxalic acid be classified as according to its ability to donate a number of $H^+_{(aq)}$ ions? (1 MARK)

d The reaction given in **c** actually occurs in two distinct stages. The equations for these reactions are given.

$(COOH)_{2(aq)} + OH^-_{(aq)} \rightarrow (COOH)(COO)^-_{(aq)} + H_2O_{(l)}$

$(COOH)(COO)^-_{(aq)} + OH^-_{(aq)} \rightarrow (COO)_2^{2-}{}_{(aq)} + H_2O_{(l)}$

There are three amphiprotic chemical species in these reactions. Give the molecular formulas of these amphiprotic species, their conjugate acids and their conjugate bases. (2 MARKS)

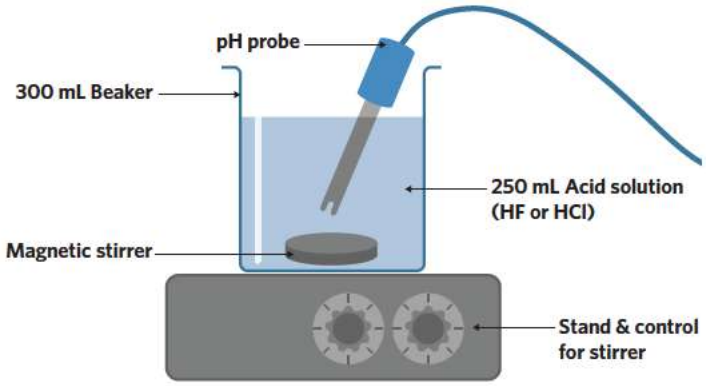
Key science skills

Question 20 (15 MARKS)

Caoimhe is a chemistry student interested in investigating the relative strengths of hydrofluoric acid, HF, and hydrochloric acid, HCl. Various sections of her experimental report are shown.

Aim

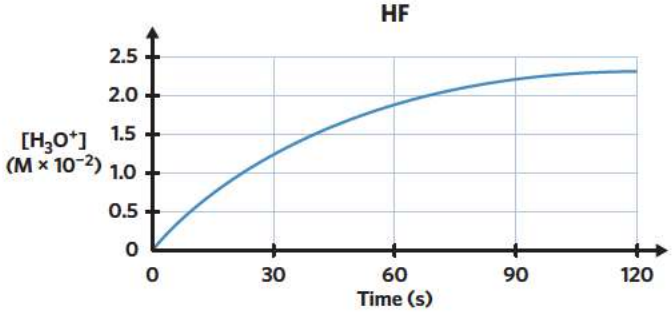
To determine the relative strengths of HF and HCl in aqueous solution by measuring their extent of ionisation over time.

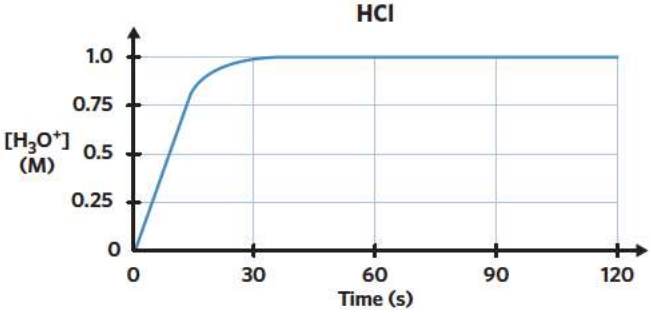


Method

1. Two 300 mL beakers were rinsed and 250 mL of deionised water was added to each.
2. A pH probe was added to each beaker to measure the concentration of $H_3O^+_{(aq)}$ ions in the solution.
3. 5.00 g of $HF_{(s)}$ was weighed in a pressurised container and then added to one of the beakers.
4. 9.125 g of $HCl_{(s)}$ was weighed in a pressurised container and then added to the second beaker.
5. The solutions were stirred using an automatic magnetic stirring bead until all solid had dissolved.
6. The results from the pH probe were recorded over 2 minutes.
7. Steps 1–6 were repeated 4 times and all results were averaged.

Averaged results





a Caoimhe wishes to ensure that she is following correct experimental procedures.

i Identify the independent and dependent variables in this experiment. (2 MARKS)

ii Identify three variables that should be controlled in this experiment. (1 MARK)

b One of Caoimhe's classmates analysing her method is concerned that she has made a number of errors.

i Why were the masses of HF and HCl used in the experiment different? Justify your answer with the use of a calculation. (2 MARKS)

ii What is the purpose of step 7 in the method? What impact, if any, does this have on the accuracy of the results obtained? (2 MARKS)

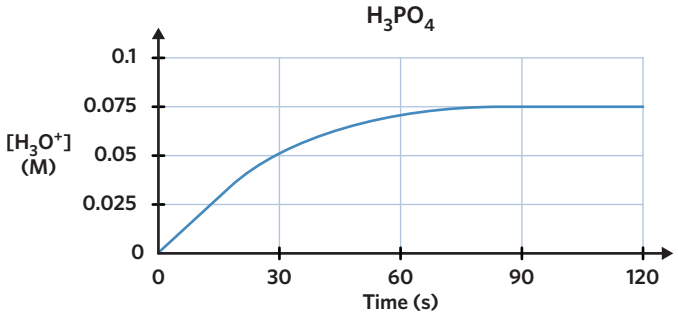
c How can Caoimhe use the data she has collected to determine each acid's strength? Is hydrofluoric acid or hydrochloric acid stronger? Justify your answer with reference to the data collected. (3 MARKS)

d While doing research about the risks associated with her experiment, Caoimhe reads the following statement:

> 'Hydrofluoric acid is corrosive to tissue due to its proton and is corrosive to bones because of its fluoride. HF also interferes with nerve function, meaning that low concentration burns may not be noticed until a significant portion of tissue and bone has been dissolved. This makes it one of the most dangerous inorganic acids. HF corrodes glass, whereas HCl does not. This is because silicon tetrafluoride exists as a more stable complex than the unstable silicon tetrachloride.'

Identify two safety measures that Caoimhe should put in place to ensure safe laboratory practice. (2 MARKS)

e Caoimhe now repeats the experiment except that this time she uses an appropriate amount of phosphoric acid. Her results are given.



Is this a valid addition to the experimental results if she intends to compare the strengths of HCl, HF and H_3PO_4? Explain. (3 MARKS)

Exam-style question hints

14: Sulfuric acid, H_2SO_4, is a common strong acid. **15:** The highly electronegative oxygen in H_2O will be attracted towards the highly positive H^+ cation in the $H_9O_4^+$ representation. **16:** "...the hydronium ion, which remains the simplest and most practical way to indicate the covalent attachment of H+ ions to water." **17a:** Acids can be classified according to how many protons they can donate. **17b:** During acid-base reactions, bases accept a proton donated by an acid. **17ci:** Conjugate bases are the species formed when an acid donates a proton. **17cii:** Conjugate acid-base pairs have opposite relative strengths. **18a:** It becomes more difficult for polyprotic acids to lose a proton after each deprotonation. **18b:** When a polyprotic acid donates a proton, it forms a negatively charged species. **19a:** Hydrogen bonding involves a H atom 'flanked' by two highly electronegative atoms. **19bi:** % composition by mass includes all species present in the solution. **19bii:** An acid's strength depends on how readily it will donate a proton. **19ci:** The number of molecules in a mole is related to avogadro's constant. **19cii:** The mass of a reactant depends on how many moles of that species are reacting. **19ciii:** Acids are classified according to the number protons they can donate. **19d:** Amphiprotic species can both donate and accept protons. **20ai:** Changing the independent variable affects the dependent variable. **20aii:** Variables are controlled so that the experimenter knows that only their independent variable is responsible for influencing the dependent variable. **20bi:** The concentration of an acidic solution will depend on how much of that acid is added. **20bii:** Random errors occur in some experiment trials, but not in all of them. **20c:** An acid's strength is related to its extent of ionisation. **20d:** Safety measures are intended to reduce the risk to whomever is performing the experiment. **20e:** The experiment is currently set up to measure only the total amount of $H_3O^+_{(aq)}$ ions in the solution.

8B MEASURING ACIDITY

In this lesson, we will investigate how basic and acidic solutions can be measured and analysed.

8A An introduction to acids and bases | 8B Measuring acidity | 8C Reactions of acids | 8D Acid-base chemistry in the body

Study design dot point

- the ionic product of water, the pH scale and the use of pH in the measurement and calculations of strengths of acids and bases and dilutions of solutions (calculations involving acidity constants are not required)

Key knowledge units

The ionic product of water	2.1.7.1
The pH scale	2.1.7.2
Effect of dilution on pH	2.1.7.3

Key terms and definitions of this lesson

Ionic product expression of the ions present in water at 25°C given by $K_W = [OH^-][H_3O^+] = 10^{-14}\ M^2$

Autoionisation process by which an atom or molecule spontaneously forms ions

Neutral solution hydronium and hydroxide concentrations are equal: $[H_3O^+] = [OH^-]$

Acidic solution hydronium is more concentrated than hydroxide: $[H_3O^+] > [OH^-]$

Basic solution hydroxide is more concentrated than hydronium: $[H_3O^+] < [OH^-]$

pH measure of the concentration of H_3O^+ ions in a solution using the formula $pH = -\log_{10}[H_3O^+]$

pH scale values of pH arranged on a logarithmic scale

Indicator substance that will experience a colour change under a specific condition

Dilution process of decreasing the concentration of a solution by adding more solvent, usually water

Lesson links

This lesson builds on:

- 8A An introduction to acids and bases

Acid-base reactions, strength and concentration form the foundation of measuring the acidity of solutions.

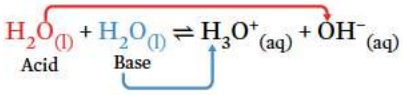
The ionic product of water 2.1.7.1

OVERVIEW

Solutions of pure water contain both hydronium and hydroxide ions, but only in low concentrations.

THEORY DETAILS

In the previous lesson, we discovered that water, as an amphiprotic substance, is capable of acting as either an acid or a base during an acid-base reaction. Therefore, although water is both a very weak base and a very weak acid, it can nonetheless accept and donate protons. In fact, two water molecules are actually capable of reacting with each other (one as the base and the other as the acid), as shown by the given equation.



This process by which water molecules spontaneously gain an electric charge is known as **autoionisation**. Of course, since water is such a weak acid and base, this reaction only occurs to a very small extent. This means that in pure water, there will only be very low concentrations of $H_3O^+_{(aq)}$ and $OH^-_{(aq)}$. In fact, for pure water at a temperature of 25°C, those concentrations are both equal to 1.00×10^{-7} M. That is:

$$[H_3O^+_{(aq)}] = [OH^-_{(aq)}] = 1.00 \times 10^{-7}\ M$$

To try and imagine how low this concentration is (0.0000001 M), consider an olympic-sized swimming pool as shown in figure 1. If it were a 2 500 000 L pool of pure water, $H_2O_{(l)}$, then the volumes of $H_3O^+_{(aq)}$ and $OH^-_{(aq)}$ would each be approximately equal to a 5 mL teaspoon.

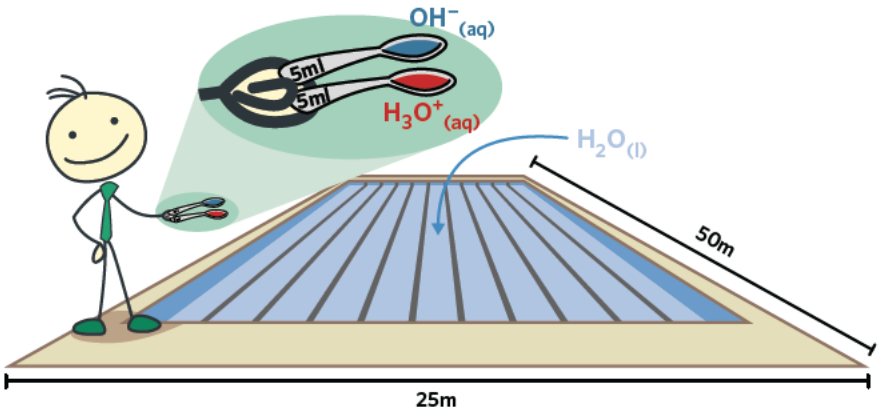


Figure 1 The concentrations of $H_3O^+_{(aq)}$ and $OH^-_{(aq)}$ in pure water are incredibly low.

What is special about these values of hydroxide and hydronium ion concentrations is that their product is always equal to 1.00×10^{-14} M^2 at 25°C. This value is known as the **ionic product** of water and is represented by the symbol K_W.

$K_W = [H_3O^+] \times [OH^-] = 1.00 \times 10^{-7}$ M $\times$ 1.00×10^{-7} M = 1.00×10^{-14} M^2

This means that if the concentration of H_3O^+ increases, the concentration of OH^- must adjust and will decrease to keep their product constant. Likewise if the concentration of H_3O^+ decreases, the concentration of OH^- must increase so that the product of their concentrations remains equal to 1.00×10^{-14} M^2. This balancing act is demonstrated in figure 2. Note that as long as the solution remains at 25°C, K_W will not change.

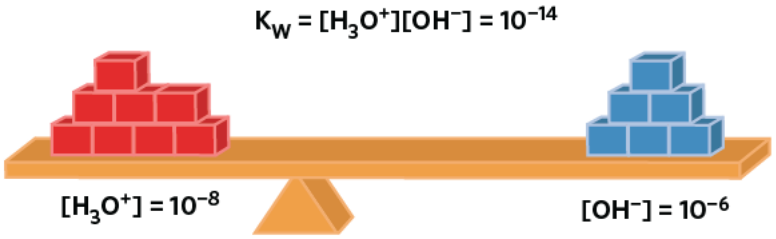


Figure 2 At 25°C the concentrations of $H_3O^+_{(aq)}$ and $OH^-_{(aq)}$ in pure water will adjust themselves so that the ionic product of water, K_W, remains equal to 1.00×10^{-14} M^2.



Note that the units of M (mol L^{-1}) when multiplied together (M × M) become the units of M^2 or (mol L^{-1})^2 = mol^2 L^{-2}. Although it is difficult to make any real physical sense of a unit like M^2 or mol^2 L^{-2}, these units are still included by convention.



Remember that the relationship $K_W = [H_3O^+] \times [OH^-] = 1.00 \times 10^{-14}$ M^2 is only true if the solution is at a temperature of 25°C.

Worked example 1

The concentration of hydronium ions in an aqueous solution at 25°C is equal to 5.32×10^3 M. Calculate the concentration of hydroxide ions present in this solution.

What information is presented in the question?

The solution is aqueous, containing water.

The concentration of hydronium ions, $[H_3O^+]$.

The temperature of the solution.

What is the question asking us to do?

Calculate the concentration of hydroxide ions, $[OH^-]$.

What strategy(ies) do we need in order to answer the question?

1. Use the relationship $K_W = [H_3O^+] \times [OH^-] = 1.00 \times 10^{-14}$ M^2 at 25°C.

Answer

Since the water is at 25°C, the relationship $K_W = [H_3O^+][OH^-] = 10^{-14}$ M^2 holds.

We are given that $[H_3O^+] = 5.32 \times 10^3$ M.

$K_W = [H_3O^+][OH^-] = 10^{-14}$ M^2

$5.32 \times 10^3 \times [OH^-] = 10^{-14}$

$$[OH^-] = \frac{10^{-14}}{5.32 \times 10^3}$$

$[OH^-] = 1.88 \times 10^{-18}$ M

The values of $[H_3O^+]$ and $[OH^-]$ in pure water, although very small, provide a baseline for a **neutral solution**. A neutral solution is one in which the concentrations of $[H_3O^+]$ and $[OH^-]$ are equal. As we have already learned, both these concentrations are equal to 10^{-7} M at 25°C.

If we add an acid to a solution of pure water, then it will form new $H_3O^+_{(aq)}$ ions by donating protons to $H_2O_{(l)}$ molecules. Since the ionic product of water remains constant at 10^{-14} M^2, this increase in $[H_3O^+]$ above the baseline of 10^{-7} M results in a decrease of $[OH^-]$. As a result, an **acidic solution** at 25°C is one in which $[H_3O^+] > 10^{-7}$ M and $[OH^-] < 10^{-7}$ M.

Similarly, adding a base to pure water results in the formation of $OH^-_{(aq)}$ ions. This increase in $[OH^-]$ will result in the decrease of $[H_3O^+]$ so that both values move away from the baseline at 10^{-7} M. Therefore, we define a **basic solution** at 25°C as one in which $[OH^-] > 10^{-7}$ M and $[H_3O^+] < 10^{-7}$ M.

In summary, the higher the concentration of $[H_3O^+]$, the more acidic a solution is. The higher the concentration of $[OH^-]$, the more basic a solution is. This relationship is shown in figure 3.

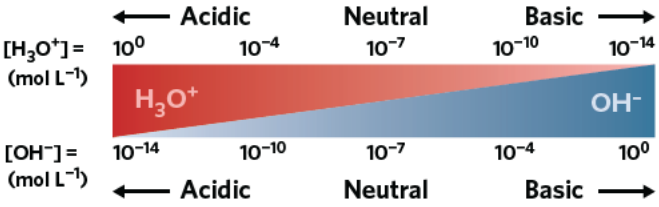


Figure 3 The acidity of a solution depends on the relative concentrations of H_3O^+ and OH^- ions. At 25°C, the product of these two concentrations will be equal to 10^{-14} M^2.

The pH scale 2.1.7.2

OVERVIEW

The pH scale is used to quantitatively measure how acidic or basic a solution is.

THEORY DETAILS

In the previous lesson, we discovered that when an acid is added to water, it results in the production of $H_3O^+_{(aq)}$ ions. If a base is added to water, it results in the production of $OH^-_{(aq)}$ ions. Therefore, we can use the concentrations of $H_3O^+_{(aq)}$ and $OH^-_{(aq)}$ to determine just how acidic or basic the solution we are investigating is.

However, in the previous section of this lesson, we also discovered that at 25°C, the concentrations of $H_3O^+_{(aq)}$ and $OH^-_{(aq)}$ are related by the equation $K_W = [OH^-][H_3O^+] = 10^{-14}$ M^2. This means we only need to know the concentration of one of these ions to determine the concentration of the other. By convention in chemistry we primarily use the concentration of H_3O^+ to measure the acidity or basicity of a solution. This measure of acidity is known as the pH of the solution and can be calculated from the following formula:

$$pH = -\log_{10}[H_3O^+]$$

or, rearranged:

$$[H_3O^+] = 10^{-pH}$$

pH values are usually considered in the range of 0–14, though incredibly strong solutions can exceed these limits. Examples of acids and bases and where their solutions fit on the pH scale are given in figure 4. Since the expression for pH involves a logarithm ($\log_{10}$), every increase or decrease in pH by 1 means that there was a change in the concentration of H_3O^+ by a factor of 10. Similarly, if there is a change in pH of 4, then the concentration of H_3O^+ has been changed by a factor of 10 000! This allows a very wide range of H_3O^+ concentrations to be represented on a relatively small scale.

We can use pH values to define what is considered an acidic, basic or neutral solution at 25°C:

- pH of 7 (pure water) is a neutral solution.
- pH less than 7 is an acidic solution.
- pH greater than 7 is a basic solution.

pH is usually considered on a scale from 0–14, with 0 representing a highly acidic solution and 14 representing a highly alkaline solution. However, if either $H_3O^+_{(aq)}$ or $OH^-_{(aq)}$ are present in high enough concentrations, pH values below 0 or above 14 are possible.



Some scientific calculators will need you to manually put the base number '10' into the $\log_{10}$ expression. Others will just have the expression log that has a default base of 10. It's important to familiarise yourself with how to use your scientific calculator for chemistry.



Since pH values depend on the negative logarithm of $[H_3O^+]$, the pH of a solution will *decrease* as the concentration of H_3O^+ ions *increases*.

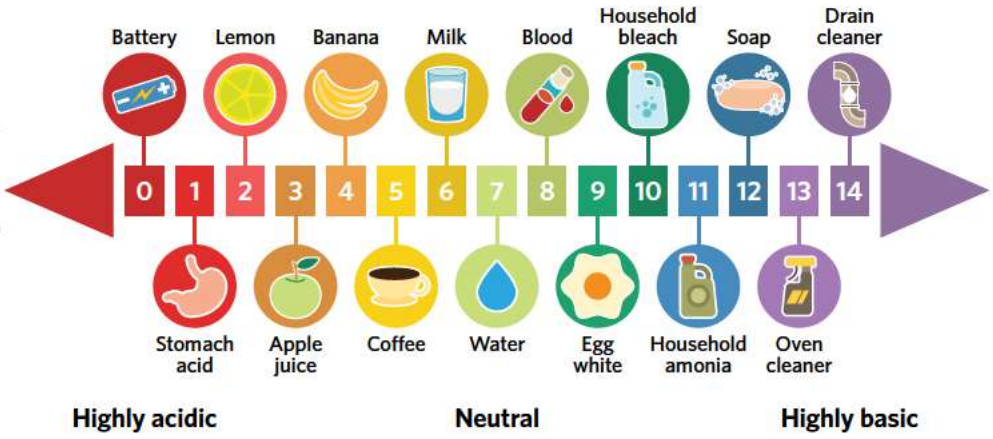


Image: VectorMine/Shutterstock.com

Figure 4 Examples of acidic and basic solutions with their corresponding pH.

In 8A, we classified acids and bases according to whether they were strong or weak. This is a qualitative description. The pH scale gives us a way to quantitatively measure how many $H_3O^+_{(aq)}$ ions are present in an acidic or basic solution.



When performing calculations to determine pH, note that the number of *decimal places* in your final answer will be equal to the number of significant figures in the given concentrations or amounts. For example, a pH of 5.3 corresponds to a concentration known to one significant figure. A pH of 5.33 corresponds to a concentration known to two significant figures.

Worked example 2

Calculate the pH of a 0.30 M solution of hydrochloric acid.

What information is presented in the question?

The concentration of a hydrochloric acid solution.

What is the question asking us to do?

Determine the solution's pH.

What strategy(ies) do we need in order to answer the question?

1. Determine whether HCl will ionise completely.
2. Determine the concentration of $H_3O^+_{(aq)}$ in the solution.
3. Use the formula $pH = -\log_{10}[H_3O^+]$.

Answer

HCl is a strong acid that will ionise completely according to the given equation.

$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

As a result, the concentration of $H_3O^+_{(aq)}$ will be equal to the concentration of $HCl_{(aq)}$.

$[HCl] = [H_3O^+] = 0.30\ M$

$pH = -\log_{10}[H_3O^+]$

$pH = -\log_{10}[0.30]$

$pH = 0.52$

Worked example 3

Calculate the pH of the following solutions:

a. 5.20 g of $NaOH_{(s)}$ added to a 7.00 L water solution

b. 3.20 M solution of barium hydroxide, $Ba(OH)_{2(aq)}$

a. 5.20 g of $NaOH_{(s)}$ added to a 7.00 L water solution

What information is presented in the question?

The mass of solid NaOH present in the solution.

The volume of the solution.

What is the question asking us to do?

Determine the pH of the NaOH solution.

What strategy(ies) do we need in order to answer the question?

1. Determine the number of mol of NaOH in the solution.
2. Consider the dissociation reaction of NaOH in water.
3. Calculate the concentration of $OH^-_{(aq)}$ ions in the solution.
4. Use the ionic product of water, K_W, to calculate the concentration of $H_3O^+_{(aq)}$ ions in the solution.
5. Use the formula $pH = -\log_{10}[H_3O^+]$.

Answer

As a strong base, sodium hydroxide dissociates completely according to the given equation.

$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

$n(OH^-_{(aq)}) = n(NaOH_{(s)}) = \frac{m}{M} = \frac{5.20}{40.0} = 0.130 \text{ mol}$

$[OH^-] = \frac{n}{V} = \frac{0.130}{7.00} = 0.01857 \text{ M}$

$K_W = [H_3O^+] \times [OH^-] = 10^{-14} \text{ M}^2$

$[H_3O^+] \times 0.01857 = 10^{-14}$

$[H_3O^+] = \frac{10^{-14}}{0.01857} = 5.385 \times 10^{-13}$

$pH = -\log_{10}[H_3O^+]$

$pH = -\log_{10}[5.385 \times 10^{-13}]$

$pH = 12.269$

---

b. 3.20 M solution of barium hydroxide, $Ba(OH)_{2(aq)}$

What information is presented in the question?

The concentration of a $Ba(OH)_{2(aq)}$ solution.

What is the question asking us to do?

Determine the pH of the $Ba(OH)_{2(aq)}$ solution.

What strategy(ies) do we need in order to answer the question?

1. Consider the dissociation reaction of $Ba(OH)_2$ in water.
2. Calculate the concentration of $OH^-_{(aq)}$ ions in the solution.
3. Use the ionic product of water, K_W, to calculate the concentration of $H_3O^+_{(aq)}$ ions in the solution.
4. Use the formula $pH = -\log_{10}[H_3O^+]$.

Answer

As a strong base, barium hydroxide dissociates completely according to the given equation.

$Ba(OH)_{2(aq)} \rightarrow Ba^{2+}_{(aq)} + 2OH^-_{(aq)}$

From this reaction, for every one mole of $Ba(OH)_2$, two moles of OH^- are produced.

$[OH^-] = 2 \times [Ba(OH)_2]$

$[OH^-] = 2 \times 3.20 = 6.40 \text{ M}$

$K_W = [H_3O^+] \times [OH^-] = 10^{-14} \text{ M}^2$

$[H_3O^+] \times 6.40 = 10^{-14}$

$[H_3O^+] = \frac{10^{-14}}{6.40} = 1.5625 \times 10^{-15}$

$pH = -\log_{10}[H_3O^+]$

$pH = -\log_{10}[1.5625 \times 10^{-15}]$

$pH = 14.806$

For the purposes of VCE chemistry, we will only investigate how to calculate the pH of a strong acid or base solution using only their initial concentration. This is because their essentially complete ionisation in water gives much simpler calculations. Calculations involving weak acids or bases and their pH are restricted to qualitative observations.

For example, it is important to note that a solution of 1.0 M hydrochloric acid (strong acid) will have a much lower pH than a solution of 1.0 M ethanoic acid (weak acid) due to its much greater degree of ionisation, resulting a higher concentration of H_3O^+ ions. A 1.0 M solution of HCl will have a pH of 0.00, while a 1.0 M solution of ethanoic acid will have a pH of 3.30.



Strong and weak acids or bases may have the same concentration (number of mol of that species per litre) but will have very different pHs due to the number of H_3O^+ ions that are formed.

Indicators

There are a number of ways to measure the pH of a solution. A standard method of measuring how acidic or basic a solution is uses the acid-base properties of plant extracts or other compounds known as **indicators**. Indicators interact with an acid or base present in a solution and will undergo a colour change in a given pH range.

For example, one of the simplest indicators is litmus paper, which uses a dye derived from lichen. As shown in figure 5, litmus paper will turn red in an acidic solution and will turn blue in a basic solution. Indicators tend to be themselves acids or bases, which allows them to interact with OH^- and H_3O^+ ions in solution and undergo a chemical change which results in an observable colour change.

Figure 6 shows a number of commonly used indicators and the colour changes they will experience in different pH ranges. Note that many indicators can have multiple colour transitions in varying pH ranges. Universal indicator is formed from the combination of many different indicators giving it its wide spectrum of colours. For this reason, the colour of universal indicator (UI) can be used to determine the approximate pH of an acidic or basic solution.

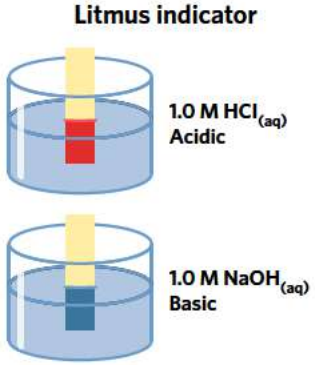


Figure 5 Acidic solutions will turn litmus paper red, whereas basic solutions will result in a blue colour.

Indicator	Colour change (0 1 2 3 4 5 6 7 8 9 10 11 12 13 14)	pH range(s)
Thymol blue	red → yellow → blue	1.2–2.8, 8.0–9.6
Methyl orange	red → orange	3.1–4.4
Bromophenol blue	yellow → blue	3.0–4.6
Methyl red	red → yellow	4.4–6.2
Bromothymol blue	yellow → blue	6.0–7.6
Phenol red	yellow → red	6.8–8.4
Phenolphthalein	colourless → pink	8.3–10.0
Universal indicator		0.0–14.0

Figure 6 Common indicators, their colour changes and pH ranges.

It's worth noting that these indicators can be found in the Year 12 data book; a resource that can be used during the VCAA Chemistry exam.

Effect of dilution on pH 2.1.7.3

OVERVIEW

Diluting an acidic or basic solution will shift its pH closer to a neutral pH of 7.

THEORY DETAILS

We have so far established that pH is a measure of the concentration of H_3O^+ ions present in a solution. We have also shown that an increase in the concentration of H_3O^+ ions will result in a decrease in the measured pH. We can increase the number of H_3O^+ ions in a solution by adding an acid to a solution of water. But what happens when we add more water to an acidic or basic solution?

By adding water we decrease the concentration of H_3O^+ ions in a solution. This addition of water is known as a **dilution**. The number of mol of H_3O^+ or OH^- ions has not changed in a strongly acidic or basic solution, but their concentrations have both decreased due to the increase in volume. This process is shown in figure 7.



Remember that since during a dilution the number of mol, *n*, of a species in solution does not change, the formula $c_1V_1 = c_2V_2$ can be used.

c_1V_1 refers to the concentration (mol L^{-1}) and volume (L) of the initial solution. c_2V_2 refers to the concentration and volume of the solution after dilution.

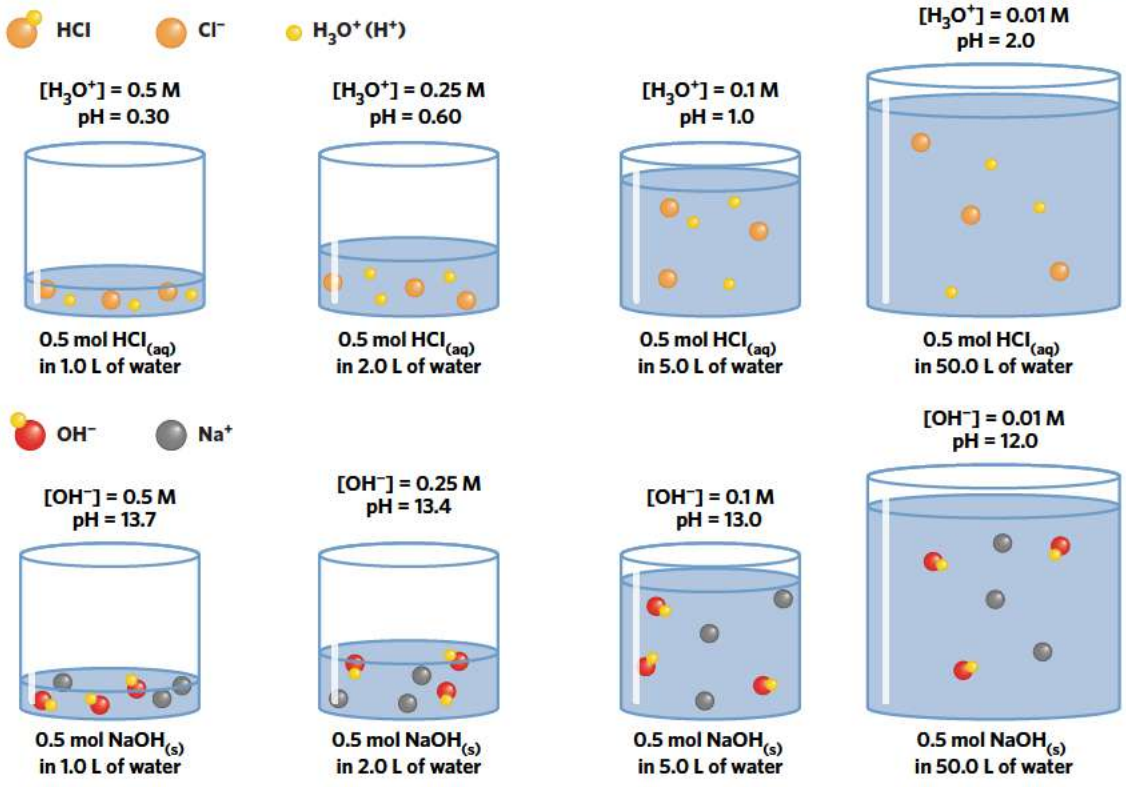


Figure 7 Increasing the volume of a solution by adding water will bring its pH closer to 7 for both acidic and basic solutions. Note that since pH is a logarithmic scale, it will take a very large volume of water to eventually reach a pH of 7.

As we can see in figure 7, as both the acidic and basic solutions become more dilute, their pHs increase or decrease. If we continued diluting these solutions, we would find that the pH would approach 7 over time. This means that all dilutions result in the pH approaching neutral (a pH value of 7), whether the original solution was classified as acidic or basic. This would make sense as we are essentially diluting the solution until the concentrations of H_3O^+ and OH^- ions are so small that they are negligible, just like in a pure water solution. Also note that, assuming that the solution is still at 25°C, the value of K_W will not have changed either and can still be used to calculate the concentrations of both H_3O^+ and OH^-.

We will only consider the dilutions of strong acids and bases, since we can assume that they undergo complete ionisation. This means that the number of H_3O^+ and OH^- will not change as we dilute the solution and our pH calculations are still accurate. The effect on pH for dilutions of weak acids and bases is beyond the scope of VCE chemistry.

Worked example 4

If 30.0 mL of water is added to a 90.0 mL 2.70 M solution of nitric acid, $HNO_{3(aq)}$, calculate the concentration of the diluted solution.

What information is presented in the question?

The volume of the initial solution and the volume of water added.

The concentration of the initial solution.

What is the question asking us to do?

Calculate the concentration of the $HNO_{3(aq)}$ solution after the dilution.

What strategy(ies) do we need in order to answer the question?

1. Calculate the new volume of the diluted solution.
2. Use the formula $c_1V_1 = c_2V_2$ to calculate the unknown concentration.

Answer

Nitric acid is a strong acid and so we can use the formula $c_1V_1 = c_2V_2$ for this dilution.

The volume of the diluted solution:

$90.0 + 30.0 = 120.0$ mL

$c_1V_1 = c_2V_2$

$$2.70 \times \frac{90.0}{1000} = c_2 \times \frac{120.0}{1000}$$

$$c_2 = \frac{2.70 \times 0.0900}{0.1200} = 2.03 \text{ M}$$

Worked example 5

Calculate the pH of a 400.0 mL solution of 0.050 M hydrochloric acid, HCl, after an extra 5.0 L of water is added.

What information is presented in the question?

The volume of the initial solution and the volume of water added.

The concentration of the initial solution.

What is the question asking us to do?

Calculate the pH of the HCl solution after the dilution.

What strategy(ies) do we need in order to answer the question?

1. Calculate the new volume of the diluted solution.
2. Use the formula $c_1V_1 = c_2V_2$ to calculate the unknown concentration.
3. Determine the concentration of $H_3O^+_{(aq)}$ in the solution.
4. Use the formula $pH = -\log_{10}[H_3O^+]$.

Answer

Hydrochloric acid is a strong acid and so we can use the formula $c_1V_1 = c_2V_2$ for this dilution. The volume of the diluted solution:

$0.4000 + 5.0 = 5.4$ L

$c_1V_1 = c_2V_2$

$0.050 \times 0.4000 = c_2 \times 5.4$

$$[HCl] = c_2 = \frac{0.050 \times 0.4000}{5.4} = 0.0037 \text{ M}$$

HCl ionises completely according to the given equation.

$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

$[HCl_{(aq)}] = [H_3O^+_{(aq)}] = 0.0037$ M

$pH = -\log_{10}[H_3O^+]$

$pH = -\log_{10}(0.0037) = 2.43$

When performing pH calculations involving dilutions of bases, it is important to first calculate the concentration of OH^- ions in the solution. It is the concentration of OH^- ions that is diluted, with the concentration of H_3O^+ ions adjusting in order to maintain the expression for the ionic product of water, $K_W = [H_3O^+] \times [OH^-] = 10^{-14}\ M^2$. Therefore, in a basic solution the concentration of OH^- ions determines the pH, not the comparatively much smaller (if not negligible) concentration of H_3O^+ ions.

Worked example 6

The pH of a 7.0 L solution of sodium hydroxide, NaOH, is equal to 12.00. An additional 550 L of water is then added to the solution.

a. Calculate the concentration of H_3O^+ ions present in the diluted solution.

b. Determine the volume of water that would need to be added to reduce the already diluted solution's pH to 8.00.

---

a. Calculate the concentration of H_3O^+ ions present in the diluted solution.

What information is presented in the question?

The volume of the initial solution and the volume of water added.

The pH of the initial solution.

What is the question asking us to do?

Determine the concentration of H_3O^+ ions in the diluted solution.

What strategy(ies) do we need in order to answer the question?

1. Use the formula $[H_3O^+] = 10^{-pH}$ to calculate the initial concentration of H_3O^+.
2. Use the ionic product of water, K_W, to calculate the concentration of $OH^-_{(aq)}$ ions in the initial solution.
3. Calculate the volume of the diluted solution.
4. Use the formula $c_1V_1 = c_2V_2$ to calculate the OH^- concentration of the diluted solution.
5. Use the ionic product of water, K_W, to calculate the concentration of $H_3O^+_{(aq)}$ ions in the diluted solution.

Answer

$[H_3O^+] = 10^{-pH} = 10^{-12.00}\ M$

$K_W = [H_3O^+] \times [OH^-] = 10^{-14}\ M^2$

$$[OH^-] = \frac{10^{-14}}{[H_3O^+]} = \frac{10^{-14}}{10^{-12.00}} = 0.0100\ M$$

The volume of the diluted solution:

$7.0 + 550 = 557\ L$

$c_1V_1 = c_2V_2$

$0.0100 \times 7.0 = c_2 \times 557$

$$[OH^-]_{diluted} = c_2 = \frac{0.0100 \times 7.0}{557} = 1.257 \times 10^{-4}\ M$$

$$[H_3O^+] = \frac{10^{-14}}{[OH^-]} = \frac{10^{-14}}{1.257 \times 10^{-4}} = 8.0 \times 10^{-11}\ M$$

---

b. Determine the volume of water that would need to be added to reduce the already diluted solution's pH to 8.00.

What information is presented in the question?

The desired pH through dilution.

What is the question asking us to do?

Calculate how much water would need to be added to the diluted solution to decrease its pH to 8.00.

What strategy(ies) do we need in order to answer the question?

1. Determine the concentration of OH^- that would be present in a solution with pH 8.00 using the formula $[H_3O^+] = 10^{-pH}$ and the ionic product of water, K_W.
2. Use the formula $c_1V_1 = c_2V_2$ to calculate the volume of the new diluted solution.
3. Subtract the initial volume to determine how much water was added.

Answer

$[H_3O^+] = 10^{-pH} = 10^{-8.0}$

$K_W = [H_3O^+] \times [OH^-] = 10^{-14}\ M^2$

$$[OH^-] = \frac{10^{-14}}{[H_3O^+]} = \frac{10^{-14}}{10^{-8.00}} = 1.0 \times 10^{-6}\ M$$

$[OH^-]_{diluted} = 1.257 \times 10^{-4}\ M$

$c_1V_1 = c_2V_2$

$1.257 \times 10^{-4} \times 557 = 1.0 \times 10^{-6} \times V_2$

$$V_2 = \frac{1.257 \times 10^{-4} \times 557}{1.0 \times 10^{-6}} = 70015\ L$$

$70015 - 557 = 6.9 \times 10^4\ L$

Theory summary

- The ionic product of water, K_W, is constant at 25°C and is a result of the spontaneous autoionisation of H_2O into hydronium and hydroxide ions.
  - $H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$
  - $K_W = [OH^-][H_3O^+] = 10^{-14}\,M^2$
- At 25°C:
  - Neutral solution: pH = 7, $[H_3O^+] = [OH^-] = 10^{-7}$ M
  - Acidic solution: pH < 7, $[H_3O^+] > 10^{-7}$ M and $[OH^-] < 10^{-7}$ M
  - Basic solution: pH > 7, $[H_3O^+] < 10^{-7}$ M and $[OH^-] > 10^{-7}$ M
- pH is a convenient way to measure how acidic or basic a solution is.
  - On the pH scale
    - 0 represents a highly acidic solution.
    - 7 represents a neutral solution (pure water).
    - 14 represents a highly basic solution.
  - pH is calculated using the formula $pH = -\log_{10}[H_3O^+]$
  - pH can be used to calculate the concentrations of strong acids and bases.
    - Strong and weak acids or bases may have the same concentration, but will have a different pH.
- Indicators are used to determine the pH range of a solution through colour changes.
- Diluting an acidic or basic solution through the addition of water will shift the solution's pH towards a neutral value of 7.
  - The formula $c_1V_1 = c_2V_2$ can be used for dilutions of strong acids and bases.

8B QUESTIONS

Theory review questions

Question 1

The ionic product of water, K_W, shows that a solution of pure water

A has no ions at all present.

B has hydronium and hydroxide ions, but only in very low concentrations.

Question 2

At 25°C, the ionic product of water $K_W = [H_3O^+][OH^-] = 10^{-14}\,M^2$ is

A constant.

B dependent on whether an acid or base has been added to the solution.

Question 3

An acidic solution has a _____ concentration of H_3O^+ ions and a _____ concentration of OH^- ions than a basic solution.

A higher, lower

B lower, higher

Question 4

An acidic solution has a pH _____, whereas a basic solution has a pH _____.

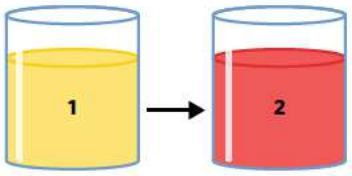
A above 7, below 7

B below 7, above 7

Question 5

The two beakers shown have a few drops of the indicator methyl orange added to them.

What is the most likely reason for this observed change?



A There was a change in the H_3O^+ concentration of the solution.

B There was a change in the overall ionic concentration of the solution.

Question 6

If the pH value of a solution decreases by 2, then the concentration of H_3O^+ ions in the solution has _____ by a factor of _____.

A decreased, 2

B increased, 100

Question 7

A solution of 1.0 M acetic acid (vinegar) has a pH of 4.0. A solution of 5.0 M acetic acid would have a _____ pH than this solution.

A lower

B higher

Question 8

Adding water to an acidic or basic solution is known as a _____.

A solvation

B dilution

Question 9

Adding an extra 600 mL of water to a 500 mL 2.0 M sulfuric acid solution will result in

A a pH decrease.

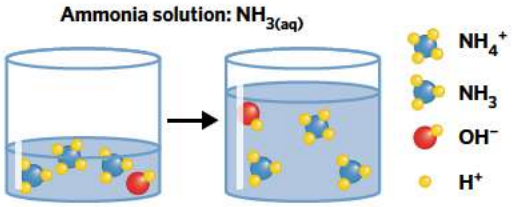
B a pH increase.

Question 10

The process shown will result in a _____ of the pH of the solution.

A decrease

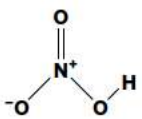
B increase



Deconstructed questions

Use the following information to answer questions 11–13.

Nitric acid, also known as 'aqua fortis' (Latin for 'strong water'), is a strong and highly corrosive acid. Nitric acid ionises in water according to the given equation and its structure is shown.



$HNO_{3(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$

A 9.3 L solution of 0.78 M $HNO_{3(aq)}$ is diluted by adding an extra 25.0 L of water to the solution. Assume that the temperature remains constant at 25°C.

Question 11 (1 MARK)

Which of the following variables will change as a result of this dilution?

I The ionic product of water, $K_W = [H_3O^+] \times [OH^-] = 10^{-14}\ M^2$

II The volume of water in the solution

III The concentration of $H_3O^+_{(aq)}$ ions in the solution

IV The concentration of $OH^-_{(aq)}$ ions in the solution

V The number of mol of $HNO_{3(aq)}$ in the solution

VI The relationship $pH = -\log_{10}[H_3O^+]$

A I, II, III, V

B II, IV, V

C II, III, IV

D I, V, VI

Question 12 (1 MARK)

Which of the following best describes how the pH of the $HNO_{3(aq)}$ solution will change as a result of the dilution?

A The pH of the solution will increase as the concentration of $H_3O^+_{(aq)}$ ions has increased.

B The pH of the solution will increase as the concentration of $H_3O^+_{(aq)}$ ions has decreased.

C The pH of the solution will decrease as the concentration of $H_3O^+_{(aq)}$ ions has increased.

D The pH of the solution will decrease as the concentration of $H_3O^+_{(aq)}$ ions has decreased.

Question 13 (2 MARKS)

Calculate the pH of the diluted solution.

Exam-style questions

Within lesson

Question 14 (1 MARK)

Which of the following correctly lists the four solutions in order of increasing pH?

A 0.01 mol L^{-1} $CH_3COOH_{(aq)}$, 2.0 M $H_2SO_{4(aq)}$, 0.025 M $NaOH_{(aq)}$, 3.7 g L^{-1} $Ba(OH)_{2(aq)}$

B 2.0 M $H_2SO_{4(aq)}$, 0.01 mol L^{-1} $CH_3COOH_{(aq)}$, 0.025 M $NaOH_{(aq)}$, 3.7 g L^{-1} $Ba(OH)_{2(aq)}$

C 0.01 mol L^{-1} $CH_3COOH_{(aq)}$, 2.0 M $H_2SO_{4(aq)}$, 3.7 g L^{-1} $Ba(OH)_{2(aq)}$, 0.025 M $NaOH_{(aq)}$

D 3.7 g L^{-1} $Ba(OH)_{2(aq)}$, 0.025 M $NaOH_{(aq)}$, 0.01 mol L^{-1} $CH_3COOH_{(aq)}$, 2.0 M $H_2SO_{4(aq)}$

Question 15 (1 MARK)

Which of the following statements is TRUE for a solution of 2.9 M $Ca(OH)_2$?

A If the solution were diluted, its pH would increase.

B Such a solution cannot exist as its pH is above the pH scale's limit of 14.

C There would be a greater concentration of OH^- ions in a 2.9 M NaOH solution than in this solution.

D In this solution, there is a lower concentration of H_3O^+ ions than OH^- ions.

Question 16 (11 MARKS)

Calculate the pH of each of the following solutions. Assume that each acid fully deprotonates in aqueous solution and that the temperature remains constant at 25°C.

a 500 mL of 0.31 M $HCl_{(aq)}$ (1 MARK)

b 30.0 L of 2.08 M $H_2SO_{4(aq)}$ added to a 12.0 L pure water solution. (3 MARKS)

c 6.4×10^2 L of 0.11 kg L^{-1} $HNO_{3(aq)}$ added to a 90.4 L pure water solution. (2 MARKS)

d In which of the above solutions would the concentration of OH^- ions be expected to be lowest? Explain your answer with reference to the ionic product of water, K_W. (2 MARKS)

e The human body requires a blood pH in the range 7.35–7.45 in order to survive. Any pH outside this range can have rapid and devastating effects on the body's ability to function.

i Is human blood acidic, basic or neutral? (1 MARK)

ii Calculate the highest and lowest concentrations of H_3O^+ ions that could be present in the blood when the human body is functioning normally. (2 MARKS)

Question 17 (12 MARKS)

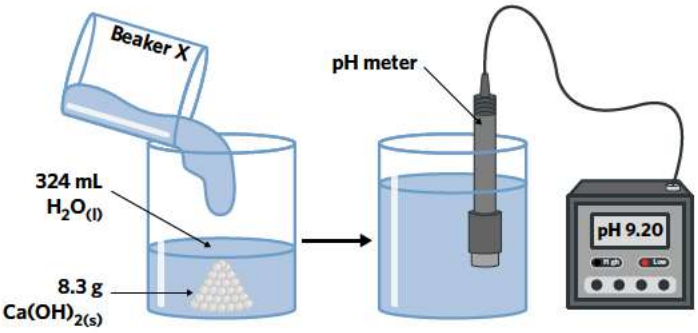
For the following questions, assume that the temperature of the solution remains constant at 25°C.

a Calculate the pH of a 70 mL 0.4 M $NaOH_{(aq)}$ solution. (1 MARK)

b Calcium hydroxide, $Ca(OH)_2$, exists as either a colourless or white crystalline solid.

i Calculate the pH of a solution of 8.3 g of $Ca(OH)_{2(s)}$ dissolved in 324 mL of pure water. (2 MARKS)

ii



How much water would need to be added to the solution from beaker X to result in the observed effect? (2 MARKS)

c Consider the following solution: 4.0 mol of $KOH_{(s)}$ is added to a 27.0 L solution of 0.90 M $KOH_{(aq)}$. The solution is then diluted by the addition of 40.4 L of water. Calculate the pH of the diluted solution. (3 MARKS)

d So far, we have only investigated the value of K_W at 25°C. In fact, the value of K_W will change depending on the temperature of the solution.

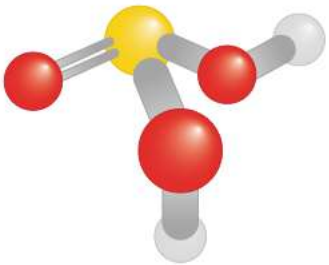
i At 100°C, the value of K_W is actually 51.3×10^{-14} M^2. How might this affect the pH value of the solution calculated in part **c**? (2 MARKS)

ii If neutral pH is still defined as when $[H_3O^+] = [OH^-]$, what would be considered neutral pH at 100°C? (2 MARKS)

Multiple lessons

Question 18 (9 MARKS)

Selenous acid, H_2SeO_3, is an acid used in the protection and alteration of the colour of steel. The 3D structure of selenous acid is shown.



Since selenous acid is not typically classified as a 'strong' acid, it does not deprotonate fully. It deprotonates according to the given two equations.

$H_2SeO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HSeO_3^-{}_{(aq)}$

$HSeO_3^-{}_{(aq)} \rightleftharpoons H^+_{(aq)} + SeO_3^{2-}{}_{(aq)}$

a What kind of acid would selenous acid be classified as? (1 MARK)

b 3.9 mol of H_2SeO_3 is added to a 72 L solution of water at 25°C.

i If 65% of H_2SeO_3 molecules deprotonate only once and 15% of H_2SeO_3 molecules deprotonate twice, calculate the number of mol of $H^+_{(aq)}$ ions present in the solution. (3 MARKS)

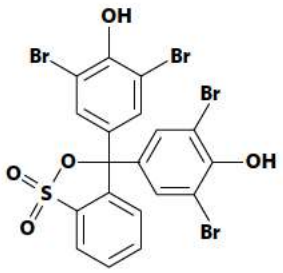
ii Calculate the pH of the solution. (1 MARK)

c Why does the second deprotonation of selenous acid happen to a much lower extent than the first deprotonation? (2 MARKS)

d If selenous acid were actually an even stronger acid, would the pH of the solution be higher or lower than the pH calculated in part **b.ii**? Explain. (2 MARKS)

Question 19 (6 MARKS)

The structure of the indicator bromophenol blue is shown. Its molecular formula is $C_{19}H_{10}Br_4O_5S$ (M = 670 g mol^{-1}). Bromophenol blue changes colour from yellow to blue as the pH of a solution transitions from 3.0 (yellow) to 4.60 (blue).



Bromophenol blue can react with ammonia, NH_3, according to the given equation.

$$C_{19}H_{10}Br_4O_5S_{(aq)} + NH_{3(aq)} \rightarrow C_{19}H_9Br_4O_5S^-_{(aq)} + NH_4^+{}_{(aq)}$$

a What would the colour of bromophenol blue be in a neutral solution? (1 MARK)

b Is bromophenol blue acting as an acid or a base in this reaction? Explain. (2 MARKS)

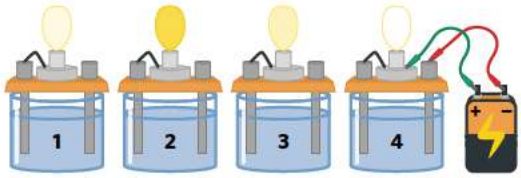
c If bromophenol blue is instead placed in 30.0 mL of a highly acidic solution with a pH of 2.00 at 25°C, what volume of water would need to be added so that the solution would completely transition to the indicator's blue colour? (2 MARKS)

d What would the concentration of $OH^-_{(aq)}$ ions in the solution be after this volume of water is added? Assume that the temperature of the solution is constant. (1 MARK)

Key science skills

Question 20 (8 MARKS)

When acids or bases are added to a solution, they form positively and negatively charged ions (H_3O^+, OH^-, etc.). These ions are capable of conducting electricity if a current is run through the solution by completing the electrical circuit. To accomplish this, an external power supply is connected to the solution.



Terry wants to investigate the electrical conductivity of acid-base solutions and so sets up the experiment as shown. He connected a battery to four different acidic solutions and measured how brightly each lamp shone. These beakers are labelled 1–4. Only one such circuit is shown in the diagram, but all lamps are shown with how brightly they shone when connected to the power supply. The results of Terry's experiment are shown in the table below.

Beaker	Solution	Brightness (lumens)
1	1.0 M $CH_3COOH_{(aq)}$	130
2	1.0 M $HNO_{3(aq)}$	780
3	1.0 M $HF_{(aq)}$	290
4	Deionised water, $H_2O_{(l)}$	0.00

Terry knows that solutions that contain high concentrations of ions are strong conductors of electricity and that solutions with low concentrations of ions are poor conductors. Assume that the temperature of all solutions remains constant at 25°C.

a What is the purpose of beaker 4 in this experiment? (1 MARK)

b Identify the independent and dependent variables in this experiment. (2 MARKS)

c List the solutions in beakers 1–4 in order of decreasing pH. Explain your reasoning. (3 MARKS)

d Terry conducts a second experiment using the solutions in the given table.

Beaker	Solution	Brightness (lumens)
1	2.0 M $CH_3COOH_{(aq)}$	
2	2.0 M $HNO_{3(aq)}$	
3	2.0 M $HF_{(aq)}$	
4	Deionised water, $H_2O_{(l)}$	

i Propose a hypothesis for the brightness of these solutions relative to those of the first experiment. (1 MARK)

ii Would the data Terry collects in this second experiment be classified as qualitative or quantitative? (1 MARK)

Exam-style question hints

14: Solutions with a lower pH have a higher concentration of H_3O^+ ions whereas solutions with a higher pH have a higher concentration of OH^- ions. **15:** In a basic solution, there is a higher concentration of OH^- ions than H_3O^+ ions. In an acidic solution, the opposite is true. **16a** The pH of a solution is related to the concentration of H_3O^+ ions. **16b:** The pH of a solution is related to the concentration of H_3O^+ ions. **16c:** The pH of a solution is related to the concentration of H_3O^+ ions. **16d:** The ionic product of water is constant and relates the concentrations of H_3O^+ and OH^-. **16e:** Whether a solution is acidic or basic is defined by its pH. **16ei:** The pH of a solution is related to the concentration of H_3O^+ ions. **17a** The pH of a basic solution is still related to the concentration of H_3O^+ ions, but the concentration of OH^- must first be calculated. **17b:** The pH of a basic solution is still related to the concentration of H_3O^+ ions, but the concentration of OH^- must first be calculated. **17bi:** Diluting a solution will bring its pH closer to 7. **17c:** pH values must take into account the total concentration of ions in a solution. **17di:** K_W is used to determine $[H_3O^+]$ and $[OH^-]$ in water. **17dii:** K_W is an expression that relates the concentrations of $[H_3O^+]$ and $[OH^-]$ to some constant value. **18a:** Acids can be classified according to how many protons they can donate. **18b:** For every deprotonation of an acid molecule, an $H^+_{(aq)}$ ion will enter the solution. **18bi:** The pH of a solution is related to the concentration of H_3O^+ (H^+) ions. **18c:** Negatively and positively charged species have a partial attractive force between them. **18d:** The strength of an acid depends on how readily it donates a proton. **19a:** Indicators change colour depending on the pH of the solution. **19b:** During an acid-base reaction, a proton is transferred from an acid to a base. **19c:** Indicators change colour when they transition through different pH ranges. **19d:** K_W is used to determine $[H_3O^+]$ and $[OH^-]$ in water. **20a** It is important to be able to make observations using a predictable baseline. **20b:** Changing the independent variable affects the dependent variable. **20c:** pH is dependent on the concentration of H_3O^+ ions present in the solution. **20di:** A hypothesis is an educated prediction of a result. **20dii:** Data can be classified according to whether it gives numerical information or describes trends.

8C REACTIONS OF ACIDS

In this lesson, we will be exploring the overall and ionic equations for reactions between acids and reactive metals, hydroxides, carbonates and hydrogen carbonates.

8A An introduction to acids and bases | 8B Measuring acidity | 8C Reactions of acids | 8D Acid-base chemistry in the body

Study design dot point

- the reactions of acids with metals, carbonates and hydroxides including balanced full and ionic equations, with states indicated

Key knowledge units

Writing ionic equations for acid-base reactions	2.1.9.1
Reactions of acids	2.1.9.2

Key terms and definitions

Spectator ion aqueous ions which do not react and remain dissolved in solution throughout a reaction

Ionic equation equation of a chemical reaction that does not include spectator ions

Reactive metal metal that will readily react with the H^+ ions from an acid

Ionic salt ionic compound formed from an anion and cation

Metal hydroxide compound of a metal cation and the hydroxide ion, OH^-

Metal carbonate compound of a metal cation and the carbonate ion, CO_3^{2-}

Metal hydrogen carbonate compound of a metal cation and a bicarbonate ion, HCO_3^-

Neutralisation reaction reaction between an acid and a metal hydroxide to produce water

Lesson links

This lesson builds on:

- 7E Solubility

Ionic compound solubility will affect ionic equations for acid-base reactions.

- 7F Precipitation reactions

Ionic equations can be written for both precipitation and acid-base reactions.

- 8B Measuring acidity

Neutralisation reactions will move a solution's pH closer to neutral pH.

Writing ionic equations for acid-base reactions 2.1.9.1

OVERVIEW

Ionic equations are a method of representing only the species that are taking part in the reaction and so are useful in exploring the reactions between acids and metal compounds.

THEORY DETAILS

In this chapter, we have so far only considered reactions between acids and bases. We will now investigate the reactions between acids and metal compounds. But first we will explore how to write **ionic equations** for chemical reactions involving acids.

Consider the given precipitation reaction between silver sulfate, $Ag_2SO_{4(aq)}$, and sodium chloride, $NaCl_{(aq)}$, in aqueous solution.

$$Ag_2SO_{4(aq)} + 2NaCl_{(aq)} \rightarrow 2AgCl_{(s)} + Na_2SO_{4(aq)}$$

As we learned in chapter 7, the compounds in the chemical equation above can be rewritten to more clearly show how all of the individual ions are reacting.

$$2Ag^+_{(aq)} + SO_4{}^{2-}_{(aq)} + 2Na^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow 2AgCl_{(s)} + 2Na^+_{(aq)} + SO_4{}^{2-}_{(aq)}$$

When the reaction is written in this way, we can see that the ions $SO_4{}^{2-}_{(aq)}$ and $Na^+_{(aq)}$ are **spectator ions**. As a result, we can simply remove them from both sides of the equation and be left with the following equation.

$$2Ag^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow 2AgCl_{(s)}$$

This can be simplified dividing both sides of the equation by 2, leaving the final equation:

$$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$$

In this lesson we will investigate writing both full and ionic equations for reactions involving acids and metal compounds.



Ionic equations do not include ions that are dissolved in the solution and are present as aqueous ions both on the reactants and products side of the reaction.

When we write ionic equations for reactions involving acids and metal compounds, it is important to remember a few key rules:

- Ionic equations, like overall equations, must still be balanced with respect to electric charge and the number of atoms of each element.
- Reactants present as solids, liquids or gases are not ions and so must be included in an ionic equation.
- Water molecules do not ionise to a significant extent and so must be present as $H_2O_{(l)}$ in an ionic equation.
- Strong acids such as nitric acid, HNO_3, ionise completely in solution to form aqueous ions such as $H^+_{(aq)}$/$H_3O^+_{(aq)}$ and $NO_3^-{}_{(aq)}$.
- $H_3O^+_{(aq)}$ is often written as $H^+_{(aq)}$ in ionic equations for simplicity.
- Soluble metal hydroxides such as sodium hydroxide, NaOH, or calcium hydroxide, $Ca(OH)_2$, will dissociate completely in solution to form metal cations ($Na^+_{(aq)}$, $Ca^{2+}_{(aq)}$) and a number of $OH^-_{(aq)}$ ions.

Reactions of acids 2.1.9.2

OVERVIEW

Acids react with metals, hydroxides and carbonates to produce an ionic 'salt' and products including hydrogen gas, carbon dioxide gas and liquid water.

THEORY DETAILS

Acids and metals

The first reaction we will consider is the reaction between a **reactive metal** and an acid. As we learned in lesson 3C, Group 1 and 2 metals are all classified as reactive metals. Reactive metals include metals such as magnesium, iron, zinc, lead, tin or calcium. Unreactive metals include copper, silver, gold and platinum. Unreactive metals will not react with acids in solution.

When an acidic solution is added to a reactive metal, the products of the reaction are an **ionic salt** and hydrogen gas, $H_{2(g)}$, giving the following general reaction.

Acid + Reactive metal → Ionic salt + $H_{2(g)}$

For example, consider the full reaction between nitric acid and iron.

$2HNO_{3(aq)} + Fe_{(s)} \rightarrow Fe(NO_3)_{2(aq)} + H_{2(g)}$

The ionic salt will be composed of the reactive metal's cation (Fe^{2+}) and the conjugate base/anion from the acid (NO_3^-), resulting in iron (II) nitrate, $Fe(NO_3)_{2(aq)}$.

The evolved hydrogen gas is observed often by an intense fizzing or bubbling at the surface of the metal. This reaction is represented in figure 1.

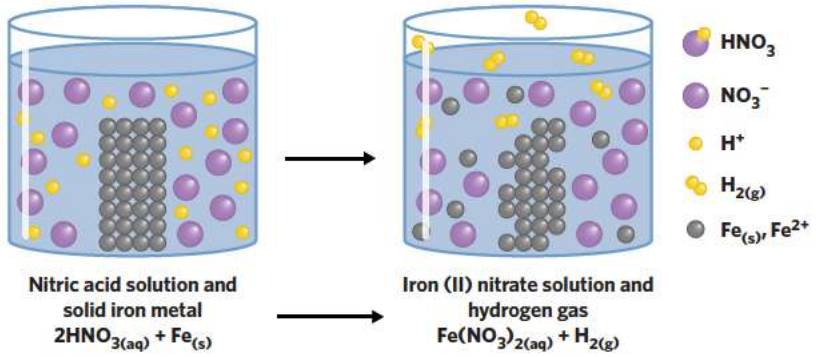


Figure 1 Nitric acid reacts with iron solid to form soluble iron (II) nitrate and hydrogen gas.

In order to determine the ionic equation for this reaction, we can rewrite the full equation as shown. Nitric acid is completely ionised into $H^+_{(aq)}$ and $NO_3^-{}_{(aq)}$ and iron (II) nitrate exists as $Fe^{2+}_{(aq)}$ and $NO_3^-{}_{(aq)}$ ions in solution.

Full equation: $2HNO_{3(aq)} + Fe_{(s)} \rightarrow Fe(NO_3)_{2(aq)} + H_{2(g)}$

Spectator ions: $2H^+_{(aq)} + 2NO_3^-{}_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + H_{2(g)}$

Ionic equation: $2H^+_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + H_{2(g)}$

Worked example 1

Write the ionic equation for the reaction between hydrochloric acid, $HCl_{(aq)}$ and solid magnesium, $Mg_{(s)}$, in aqueous solution.

What information is presented in the question?

The acid $HCl_{(aq)}$ is reacting with the reactive metal $Mg_{(s)}$ in aqueous solution (water).

What is the question asking us to do?

Write the ionic equation for this reaction.

What strategy(ies) do we need in order to answer the question?

1. Identify the general equation for an acid reacting with a reactive metal.
2. Determine which ions in the equation are spectator ions.
3. Write the ionic equation for the reaction, remembering to omit any spectator ions.

Answer

The general equation for an acid reacting with a reactive metal is:

Acid + Reactive metal → Ionic salt + $H_{2(g)}$

The ionic salt $MgCl_{2(aq)}$ will be formed from the acid's conjugate base/anion ($Cl^-_{(aq)}$) and the metal's cation ($Mg^{2+}_{(aq)}$).

Therefore, the overall balanced equation is:

$2HCl_{(aq)} + Mg_{(s)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

This equation can be rewritten to show all individual ions:

$2H^+_{(aq)} + 2Cl^-_{(aq)} + Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)} + H_{2(g)}$

Chloride, $Cl^-_{(aq)}$, appears on both sides of the equation as an ion and so is a spectator ion that can be omitted in the balanced ionic equation:

$2H^+_{(aq)} + Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)}$

If the ionic salt formed from the reaction between an acid and a reactive metal is insoluble in water, then the ionic equation will reflect this. Consider the following reaction between hydrochloric acid, $HCl_{(aq)}$, and solid lead metal, $Pb_{(s)}$.

Full equation: $2HCl_{(aq)} + Pb_{(s)} \rightarrow PbCl_{2(s)} + H_{2(g)}$

Spectator ions: $2H^+_{(aq)} + 2Cl^-_{(aq)} + Pb_{(s)} \rightarrow PbCl_{2(s)} + H_{2(g)}$

Ionic equation: $2HCl_{(aq)} + Pb_{(s)} \rightarrow PbCl_{2(s)} + H_{2(g)}$

Lead (II) chloride, $PbCl_{2(s)}$, is poorly soluble in water and so will exist mainly in solid form. Since there are no chemical species present as ions on both sides of the equation, there are no spectator ions, and the ionic equation for this reaction will be identical to its overall reaction.

As mentioned in 3C, the hydrogen gas produced from these reactions can be identified by the hydrogen pop test as shown in figure 2.

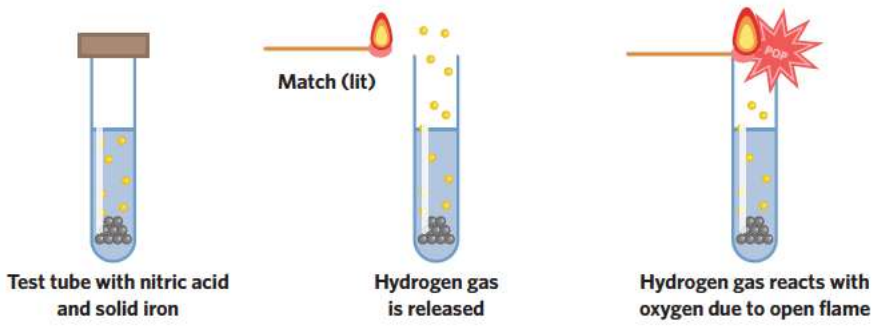


Figure 2 The presence of hydrogen gas can be tested by the hydrogen 'pop' test.

Acids and metal hydroxides

The second reaction we will consider is the reaction between a **metal hydroxide** and an acid. Metal hydroxides are ionic compounds formed from a metal cation and hydroxide ions, OH^-. Examples of metal hydroxides include sodium hydroxide, NaOH, barium hydroxide, $Ba(OH)_2$, and potassium hydroxide, KOH. When added to aqueous solution, a soluble metal hydroxide will dissociate into its cation and hydroxide ions as shown for KOH.

$KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$

When an acidic solution is added to a metal hydroxide, the products of the reaction are an **ionic salt** and liquid water, $H_2O_{(l)}$, giving the following general reaction:

Acid + Metal hydroxide → Ionic salt + $H_2O_{(l)}$

For example, consider the full reaction between potassium hydroxide, KOH, and nitric acid, HNO_3.

$HNO_{3(aq)} + KOH_{(aq)} \rightarrow KNO_{3(aq)} + H_2O_{(l)}$

Once again, the ionic salt will be composed of the metal hydroxide's cation (K^+) and the conjugate base/anion from the acid (NO_3^-), resulting in potassium nitrate, $KNO_{3(aq)}$. This reaction is represented in figure 3. Remember that in solution ionic compounds such as $KNO_{3(aq)}$ exist as aqueous forms of anions and cations. If we were to evaporate the solution's water, however, we would be left with a solid $KNO_{3(s)}$ residue.

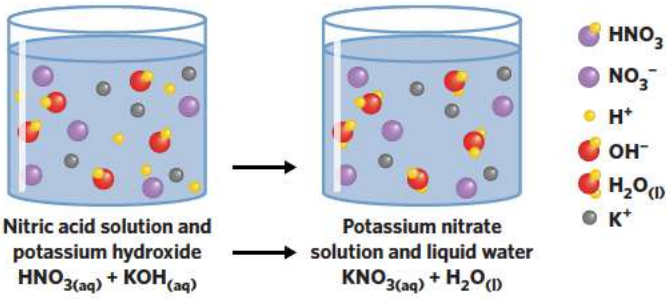


Figure 3 Nitric acid reacts with potassium hydroxide to form the soluble potassium nitrate and liquid water.

In order to determine the ionic equation for this reaction, we can rewrite the full equation as shown. Nitric acid is completely ionised into $H^+_{(aq)}$ and $NO_3^-{}_{(aq)}$ and potassium nitrate exists as $K^+_{(aq)}$ and $NO_3^-{}_{(aq)}$ ions in solution.

Full equation: $HNO_{3(aq)} + KOH_{(aq)} \rightarrow KNO_{3(aq)} + H_2O_{(l)}$

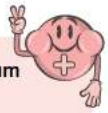
Spectator ions: $H^+_{(aq)} + NO_3^-{}_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \rightarrow K^+_{(aq)} + NO_3^-{}_{(aq)} + H_2O_{(l)}$

Since $K^+_{(aq)}$ and $NO_3^-{}_{(aq)}$ exist as ions on both sides of the equation, we can eliminate them both as spectator ions. Water is a covalent molecular compound and so must be present in the ionic equation. This leaves us with the following balanced ionic equation:

Ionic equation: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

When a reaction between an acid and a soluble metal hydroxide is shown as an ionic equation, we can see that $H^+_{(aq)}$ ions from the acid and $OH^-_{(aq)}$ ions from the base are reacting together to form water, $H_2O_{(l)}$. This is known as a **neutralisation reaction**, as the acid and base are effectively removing each other from the solution and forming water, which has a neutral pH. When all of the basic $OH^-_{(aq)}$ ions have reacted with the acidic $H^+_{(aq)}$ ions to form water, the solution is said to be 'neutralised'.

A second example of a metal hydroxide reacting with an acid is explored in worked example 2.



If the acidic hydronium ions in the given neutralisation reaction are expressed as $H_3O^+_{(aq)}$ instead of $H^+_{(aq)}$, the balanced ionic equation would be written as $H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2H_2O_{(l)}$.

Worked example 2

Write the full equation for the reaction between sulfuric acid, $H_2SO_{4(aq)}$ and sodium hydroxide, $NaOH_{(aq)}$, in aqueous solution.

What information is presented in the question?

The acid $H_2SO_{4(aq)}$ is reacting with the metal hydroxide $NaOH_{(aq)}$ in aqueous solution (water).

What is the question asking us to do?

Write the ionic equation for this reaction.

What strategy(ies) do we need in order to answer the question?

1. Identify the general equation for an acid reacting with a metal hydroxide.
2. Determine which ions will contribute to form the ionic salt.
3. Write the full (overall) equation for the reaction.
4. Balance the equation with respect to atoms and charges.

Answer

The general equation for an acid reacting with a metal hydroxide is:

Acid + Metal hydroxide → Ionic salt + $H_2O_{(l)}$

The ionic salt in this reaction will be formed from the conjugate base/anion of the acid (SO_4^{2-}) and the cation from the metal hydroxide (Na^+).

Therefore, the overall (unbalanced) equation is:

$H_2SO_{4(aq)} + NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$

The Na and H must then be balanced, leaving the overall balanced equation:

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

Acids and metal carbonates

The final reaction we will consider is the reaction between a **metal carbonate** and an acid. Metal carbonates are basic compounds formed from a metal cation and carbonate ions, CO_3^{2-}. Metal carbonates can be either soluble or insoluble in water.

Examples of soluble metal carbonates include the Group I metal carbonates: sodium carbonate, Na_2CO_3, and potassium carbonate, K_2CO_3. When added to aqueous solution, a soluble metal carbonate will dissociate to form its cation and carbonate ions, as shown for Na_2CO_3.

$Na_2CO_{3(aq)} \rightarrow 2Na^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

When an acidic solution is added to a metal carbonate, the products of the reaction are an ionic salt, carbon dioxide gas, $CO_{2(g)}$ and liquid water, $H_2O_{(l)}$, giving the following general reaction.

Acid + Metal carbonate → Ionic salt + $CO_{2(g)}$ + $H_2O_{(l)}$

For example, consider the full reaction between sodium carbonate, $Na_2CO_{3(aq)}$, and nitric acid, HNO_3.

$2HNO_{3(aq)} + Na_2CO_{3(aq)} \rightarrow 2NaNO_{3(aq)} + CO_{2(g)} + H_2O_{(l)}$

This reaction is represented in figure 4.

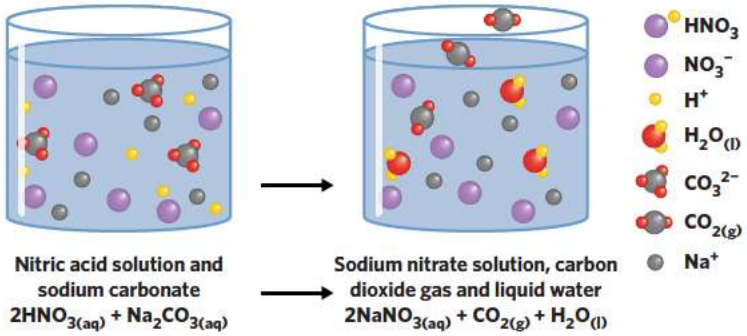


Figure 4 Nitric acid reacts with sodium carbonate to form sodium nitrate, water and carbon dioxide gas.

In order to determine the ionic equation for this reaction, we can rewrite the full equation as shown.

Full equation: $2HNO_{3(aq)} + Na_2CO_{3(aq)} \rightarrow 2NaNO_{3(aq)} + CO_{2(g)} + H_2O_{(l)}$

Spectator ions: $2H^+_{(aq)} + 2NO_3^-{}_{(aq)} + 2Na^+_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow 2Na^+_{(aq)} + 2NO_3^-{}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Water and carbon dioxide are covalent molecular compounds and so must be present in the ionic equation. This leaves us with the following balanced ionic equation:

Ionic equation: $2H^+_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

Insoluble metal carbonates also follow the same overall full reaction as soluble metal carbonates. These include magnesium carbonate, $MgCO_3$, calcium carbonate, $CaCO_3$, and zinc carbonate, $ZnCO_3$. These compounds will remain in solid form when added to aqueous solution. This means that the ionic equations for their reactions with acids will be different. An example of the reaction between nitric acid and solid calcium carbonate is given below.

Full equation: $2HNO_{3(aq)} + CaCO_{3(s)} \rightarrow Ca(NO_3)_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

Spectator ions: $2H^+_{(aq)} + 2NO_3^-{}_{(aq)} + CaCO_{3(s)} \rightarrow Ca^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Ionic equation: $2H^+_{(aq)} + CaCO_{3(s)} \rightarrow Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Closely related to metal carbonates are **metal hydrogen carbonates**, or bicarbonates. Metal hydrogen carbonates are compounds involving a metal cation and the bicarbonate ion, HCO_3^-. Examples include sodium hydrogen carbonate, $NaHCO_3$, potassium hydrogen carbonate, $KHCO_3$, magnesium hydrogen carbonate, $Mg(HCO_3)_2$ and calcium hydrogen carbonate, $Ca(HCO_3)_2$. Metal hydrogen carbonates follow the same general reaction with an acid as metal carbonates and have similar ionic equations as well. This is shown in the following example.

General equation: Acid + Metal hydrogen carbonate → Ionic salt + $CO_{2(g)}$ + $H_2O_{(l)}$

Full equation: $HNO_{3(aq)} + KHCO_{3(s)} \rightarrow KNO_{3(aq)} + CO_{2(g)} + H_2O_{(l)}$

Spectator ions: $H^+_{(aq)} + NO_3^-{}_{(aq)} + KHCO_{3(s)} \rightarrow K^+_{(aq)} + NO_3^-{}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Ionic equation: $H^+_{(aq)} + KHCO_{3(s)} \rightarrow K^+_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Worked example 3

Write the ionic equation for the reaction between hydrochloric acid, $HCl_{(aq)}$ and solid magnesium carbonate, $MgCO_{3(s)}$, in aqueous solution.

What information is presented in the question?

The acid $HCl_{(aq)}$ is reacting with the insoluble metal carbonate $MgCO_{3(s)}$ in aqueous solution (water).

What is the question asking us to do?

Write the ionic equation for this reaction.

What strategy(ies) do we need in order to answer the question?

1. Identify the general equation for an acid reacting with a metal carbonate.
2. Determine which ions in the equation are spectator ions.
3. Write the ionic equation for the reaction, remembering to omit any spectator ions.

Answer

The general equation for an acid reacting with a metal carbonate is:

Acid + Metal carbonate → Ionic salt + $CO_{2(g)}$ + $H_2O_{(l)}$

The ionic salt $MgCl_{2(aq)}$ will be formed from the acid's anion ($Cl^-_{(aq)}$) and the metal's cation ($Mg^{2+}_{(aq)}$).

Therefore, the overall balanced equation is:

$2HCl_{(aq)} + MgCO_{3(s)} \rightarrow MgCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

Note that $MgCO_3$ is insoluble in water and so will be represented as solid $MgCO_{3(s)}$. This equation can be rewritten to show all individual ions:

$2H^+_{(aq)} + 2Cl^-_{(aq)} + MgCO_{3(s)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Chloride, $Cl^-_{(aq)}$, is the only species that appears as an ion on both sides of the equation, making it a spectator ion. As a spectator ion, it can be omitted from the balanced ionic equation:

$2H^+_{(aq)} + MgCO_{3(s)} \rightarrow Mg^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Theory summary

- Ionic equations do not include chemical species that are present as aqueous ions on both sides of an equation. These species are known as spectator ions.
- The general equations for acids with different compounds:
  - Acid + Reactive metal → Ionic salt + $H_{2(g)}$
  - Acid + Metal hydroxide → Ionic salt + $H_2O_{(l)}$
  - Acid + Metal carbonate → Ionic salt + $CO_{2(g)}$ + $H_2O_{(l)}$
  - Acid + Metal hydrogen carbonate → Ionic salt + $CO_{2(g)}$ + $H_2O_{(l)}$
- These reactions can all be represented as either overall/full equations or ionic equations.

8C QUESTIONS

Theory review questions

Question 1

An ionic equation does not include species present only as aqueous ions in a chemical reaction. These species are known as _____.

A background ions

B spectator ions

Question 2

Water molecules, $H_2O_{(l)}$, are never included in ionic equations.

A True

B False

Question 3

Which of the following is the correct general equation for the reaction between an acid and a reactive metal?

A Acid + Reactive metal → Ionic salt + $H_{2(g)}$

B Acid + Reactive metal → Ionic salt + $H_2O_{(g)}$

Question 4

Which of the following is an ionic salt?

A $NaNO_3$

B CH_4

Question 5

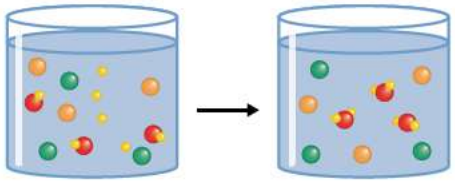
When an acid reacts with a metal hydroxide, what are the products?

A $H^+_{(aq)}$, $OH^-_{(aq)}$

B ionic salt, $H_2O_{(l)}$

Question 6

Which of the species present in this reaction are spectator ions?



A 

B 

Question 7

A neutralisation reaction is most likely to occur when an acid and a _____ react together.

A metal carbonate

B metal hydroxide

Question 8

Which of the following is a reaction between an acid and a metal carbonate?

A $2HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

B $H_3PO_{4(aq)} + 3KOH_{(aq)} \rightarrow K_3PO_{4(aq)} + 3H_2O_{(l)}$

Question 9

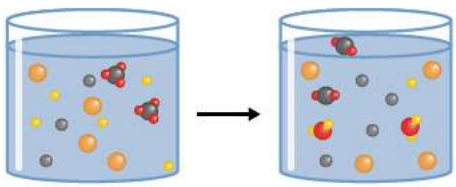
What is the ionic equation for the reaction between an acid and a metal hydroxide?

A $2HNO_{3(aq)} + Ba(OH)_{2(aq)} \rightarrow Ba(NO_3)_{2(aq)} + 2H_2O_{(l)}$

B $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Question 10

What kind of reaction is this diagram most likely representing?



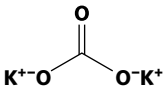
A Acid + Reactive metal

B Acid + Metal carbonate

Deconstructed questions

Use the following information to answer questions 11–13.

Potassium carbonate, K_2CO_3, is a highly soluble white salt. Its solubility is so high that over 100 g of potassium carbonate will dissolve completely in only 100 mL of water. The structure of K_2CO_3 is shown. Consider what will happen if potassium carbonate reacts with the strong acid hydroiodic acid, HI.



Question 11 (1 MARK)

What is the general equation for the reaction between hydroiodic acid and a metal carbonate?

A $HI_{(aq)}$ + Metal carbonate → Ionic salt + $CO_{2(g)}$ + $H_3O^+_{(aq)}$
B $2HI_{(aq)}$ + Metal carbonate → Ionic salt + $CO_{2(g)}$ + $H_2O_{(l)}$
C $2HI_{(aq)}$ + Metal carbonate → Cation + $CO_{2(g)}$ + $H_2O_{(l)}$
D $HI_{(aq)}$ + Metal carbonate → Anionic salt + $NO_{2(g)}$ + $OH^-_{(l)}$

Question 12 (1 MARK)

Which of the following is the correct balanced equation for the overall reaction between hydroiodic acid and potassium carbonate?

A $2HI_{(aq)} + K_2CO_{3(aq)} \rightarrow 2HK_{(aq)} + IO_{3(s)} + C_{(s)}$
B $HI_{(aq)} + K_2CO_{3(aq)} \rightarrow KI_{(aq)} + CO_{2(g)} + H_2O_{(l)}$
C $HI_{(aq)} + K_2CO_{3(s)} \rightarrow 2KI_{(aq)} + HCO_3^-{}_{(aq)}$
D $2HI_{(aq)} + K_2CO_{3(aq)} \rightarrow 2KI_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Question 13 (2 MARKS)

The reaction between an acid and a metal carbonate can also be represented as an ionic equation.

a Identify the spectator ion(s) from the overall reaction between $HI_{(aq)}$ and $K_2CO_{3(aq)}$. (1 MARK)

b Write the balanced ionic equation for the reaction between $HI_{(aq)}$ and $K_2CO_{3(aq)}$. (1 MARK)

Exam-style questions

Within lesson

Question 14 (1 MARK)

Which of the following reactions would have an ionic equation that is identical to its overall equation?

A $H_3PO_{4(aq)} + Ca_{(s)} \rightarrow Ca_3(PO_4)_{2(aq)} + H_{2(g)}$
B $CdSO_{4(aq)} + K_2S_{(aq)} \rightarrow CdS_{(s)} + K_2SO_{4(aq)}$
C $HCl_{(aq)} + NaHCO_{3(aq)} \rightarrow NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$
D $Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow 2Fe_{(s)} + Al_2O_{3(s)}$

Question 15 (1 MARK)

Which of the following is the balanced ionic equation for the reaction between solid zinc and chloric acid, $HClO_{3(aq)}$?

A $HClO_{3(aq)} + Zn_{(s)} \rightarrow Zn(ClO_3)_{2(aq)} + H_{2(g)}$
B $H^+_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$
C $2H^+_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$
D Solid zinc would not react with chloric acid.

Question 16 (5 MARKS)

Ethanoic acid, CH_3COOH, also known as acetic acid, is the acid found in household vinegar. A solution of ethanoic acid is added to a 40.0 L reaction vessel containing 2.0 M sodium hydroxide.

a Write the balanced overall equation for the reaction that would occur. (1 MARK)

b Write the balanced ionic equation for this reaction. (1 MARK)

c Explain your reasoning for the omission or inclusion of chemical species in the ionic equation in part **b**. (2 MARKS)

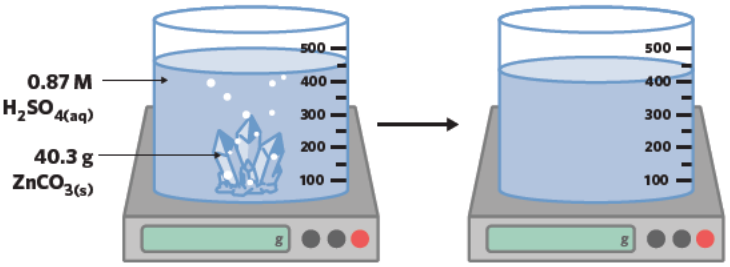
d What is the name given to this kind of reaction? (1 MARK)

Multiple lessons

Question 17 (6 MARKS)

Smithsonite is a mineral ore of zinc, with the molecular formula $ZnCO_3$.

Although insoluble in water, smithsonite can be broken down by the addition of concentrated sulfuric acid in aqueous solution. Consider the following 500 mL reaction vessel:



a Write the balanced equation for the overall reaction occurring in the beaker. (1 MARK)

b What is the balanced ionic equation for this reaction? (1 MARK)

c The chemist investigating this reaction noticed that the water level in the beaker changed after the reaction had finished. Explain this observation. (2 MARKS)

d The chemist weighed the reaction vessel before and after the reaction had finished. Calculate the change in mass that would be observed. (2 MARKS)

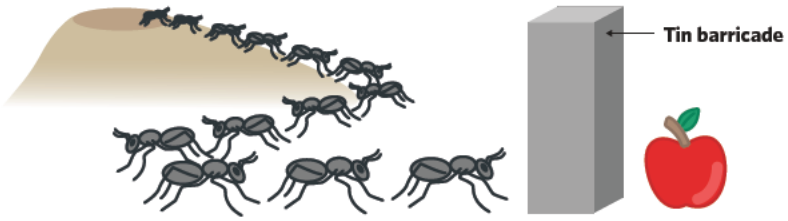
Question 18 (9 MARKS)

Methanoic acid, also known as formic acid, is used in the defense mechanism of many different ant species to protect their colonies and subdue their prey. It is a colourless monoprotic organic acid that exists in a highly soluble aqueous form in the ant's abdominal glands.

a Draw the structural formula of methanoic acid. (1 MARK)

b What is the name of the functional group present in methanoic acid? (1 MARK)

c Consider the following scenario:
A number of ants are on a scouting mission from their colony. On their journey, they encounter a barricade made out of the metal tin (Sn). These intrepid explorer ants identify the tin barricade as an enemy that is preventing them from getting to a nearby apple.



In order to defend themselves and acquire the apple, the ants spray methanoic acid onto the tin barricade to break it down.

i Write the balanced overall equation for the reaction that would occur between the methanoic acid from the ants and the tin barricade. (1 MARK)

ii Write the balanced ionic equation for this reaction. (1 MARK)

d Each ant can produce an average of 0.600 μg (1 μg = 1×10^{-6} g) of formic acid. If the tin barricade has a mass of 40.0 g, how many ants would be needed to completely break it down and reach the apple? (3 MARKS)

e How many molecules of $H_{2(g)}$ would be produced in the reaction from part **d**? (2 MARKS)

Question 19 (8 MARKS)

The human stomach contains cells that produce hydrochloric acid, HCl. The average human stomach has a volume of 1.20 L and a concentration of hydrochloric acid equal to 3.2×10^{-4} M. Although stomach tissue has evolved to handle these highly acidic conditions, the oesophagus has not. If acid leaves the stomach and comes into contact with this unprotected tissue, the result is acid reflux (heartburn).

a Calculate the pH of the average human stomach. (1 MARK)

b Is hydrochloric acid an amphiprotic, monoprotic or polyprotic acid? (1 MARK)

c Antacids are tablets taken to reduce the effects of acid reflux. A common antacid is aluminium hydroxide, $Al(OH)_3$, which comes in 250 mg tablets. Aluminium hydroxide is insoluble in water but will react with the hydrochloric acid in the stomach.

i Give the balanced overall equation for the reaction between aluminium hydroxide and hydrochloric acid. (1 MARK)

ii Give the balanced ionic equation for this reaction. (1 MARK)

d A patient suffering from a severe case of acid reflux has a stomach pH equal to 1.20. How many antacid tablets should this patient take in order to return their stomach to average acidity? Assume that the volume and temperature of the stomach remain constant. (4 MARKS)

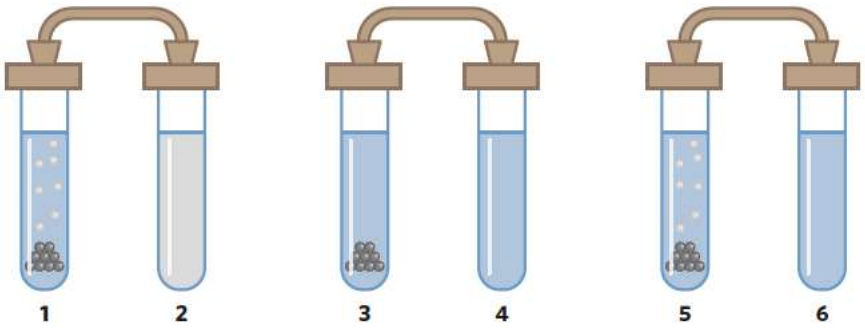
Key science skills

Question 20 (10 MARKS)

The limewater test is a test used by chemists to determine the presence of carbon dioxide gas, $CO_{2(g)}$, in a solution. Limewater is a saturated solution of calcium hydroxide, $Ca(OH)_{2(aq)}$, that will transition to a white 'milky' or 'cloudy' solution if $CO_{2(g)}$ is present. This is due to the precipitation of insoluble calcium carbonate, $CaCO_{3(s)}$, that becomes suspended in the solution.

$$Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$$

A chemist has mixed up the labels of three of her metal compounds – $Pb_{(s)}$, $MgCO_{3(s)}$ and $Zn(OH)_{2(s)}$ – and so sets up an experiment to determine their identity. To accomplish this, she prepared the three sets of test tubes shown below.



Test tube	Contents	
1	Unknown compound 1	$HNO_{3(aq)}$
2	$Ca(OH)_{2(aq)}$	
3	Unknown compound 2	$HNO_{3(aq)}$
4	$Ca(OH)_{2(aq)}$	
5	Unknown compound 3	$HNO_{3(aq)}$
6	$Ca(OH)_{2(aq)}$	

a Determine which test tube contains $MgCO_{3(s)}$. Justify your answer with reference to the data collected. (2 MARKS)

b Determine which test tubes contain $Pb_{(s)}$ and $Zn(OH)_{2(s)}$. Justify your answer with reference to the data collected. (3 MARKS)

c Are the data collected best classified as quantitative or qualitative? Explain. (2 MARKS)

d Hydrogen gas is incredibly flammable and will react violently with any heat sources. Identify two safety precautions that could be taken to ensure safe laboratory practice. (2 MARKS)

Exam-style question hints

14: Ionic equations do not include spectator ions. **15:** Zinc is a reactive metal that will react with chloric acid to produce an ionic salt from its cation (Zn^{2+}) and chloric acid's conjugate base (ClO_3^-) and hydrogen gas, $H_{2(g)}$. **16a:** Acid + Metal hydroxide → Ionic salt + $H_2O_{(l)}$ **16b:** Ionic equations do not include spectator ions. **16c:** Spectator ions are chemical species that are present as ions on both sides of a chemical equation. **16d:** Water has a neutral pH. **17a:** Acid + Metal carbonate → Ionic salt + $CO_{2(g)}$ + $H_2O_{(l)}$ **17b:** Ionic equations do not include spectator ions. **17c:** Since there is no lid on the beaker, any evolved gases will escape. **17d:** The amount of $CO_{2(g)}$ produced will depend on the amount of $ZnCO_{3(s)}$. **18a:** A structural formula shows all atoms and bonds in a molecule. **18b:** Methanoic acid contains a functional group that can donate a proton, $H^+_{(aq)}$ **18c:** Acid + Reactive metal → Ionic salt + $H_{2(g)}$ **18d:** Ionic equations do not include spectator ions. **18e:** In reactions between acids and reactive metals, the metal becomes a dissolved cation. **18f:** The amount of $H_{2(g)}$ produced will be proportional to the amount of tin consumed. **19a:** pH is related to the concentration of H_3O^+ ions in a solution. **19b:** Acids can be classified according to how many protons they can accept or receive. **19c:** Acid + Metal hydroxide → Ionic salt + water **19d:** Ionic equations do not include spectator ions. **19e:** When $H^+_{(aq)}$ ions are removed from solutions, the pH increases. **20a:** Acid + Metal carbonate → Ionic salt + $CO_{2(g)}$ + $H_2O_{(l)}$ **20b:** Acid + Reactive metal → Ionic salt + $H_{2(g)}$ **20c:** Data can be classified according to whether it gives numerical information or describes trends. **20d:** Safety measures are intended to reduce the risk to whomever is performing the experiment.

8D ACID-BASE CHEMISTRY IN THE BODY

In this lesson, we will be investigating how acids, bases and pH affect the organic molecules and functioning of the human body.

8A An introduction to acids and bases | 8B Measuring acidity | 8C Reactions of acids | 8D Acid-base chemistry in the body

Study design dot point

- the causes and effects of a selected issue related to acid-base chemistry

Key knowledge unit

Acid-base chemistry in the body	2.1.10.1

Key terms and definitions

Amino group functional group with molecular formula NH_2

Amino acid organic molecule with a central carbon attached to an amino group, a carboxyl group, a side chain (R group) and a hydrogen atom

Peptide/amide bond functional group that links amino acids together with molecular formula CONH

Protein complex organic molecules formed from multiple amino acids

Enzyme proteins capable of increasing the rate of a chemical reaction in a biological system

Activity rate at which an enzyme converts between reactants and products

Optimal pH pH at which an enzyme will be most effective

Carbonic anhydrase enzyme that increases the rate of the reaction $CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$

Hyperventilation increased breathing rate

Hypoventilation decreased breathing rate

Alkalosis health condition in which blood pH is above 7.45

Acidosis health condition in which blood pH is below 7.35

Lesson links

This lesson builds on:

- 5C Intermolecular bonding

Intermolecular forces determine the overall structure of proteins.

- 6E Functional groups - alcohols and carboxylic acids

Amino acids contain the COOH functional group.

- 7E Solubility

The solubility of gases influences the amount of gas dissolved in the bloodstream.

- 8B Measuring acidity

Blood pH alters the function of enzymes.

Acid-base chemistry in the body 2.1.10.1

OVERVIEW

Molecules with acid-base properties play important roles in maintaining a healthy human body and allowing cells to perform their vital, life-sustaining functions.

THEORY DETAILS

Biochemistry is the study of the intersection between chemistry and biological systems. Our bodies are tightly regulated by the interaction of these biological systems and, at its core, depends on trillions of underlying chemical reactions every second.

Amino acids

One of the most important molecules in biochemistry is the **amino acid**. Amino acids are organic molecules that contain both an **amino group** ($-NH_2$) and a carboxyl group ($-COOH$) attached to a central carbon, as shown in figure 1. The amino functional group is capable of accepting a proton, H^+, to form $-NH_3^+$ and the carboxyl group is capable of donating a proton to form $-COO^-$. Since amino acids exhibit the properties of both Brønsted-Lowry acids and bases, this makes them amphiprotic substances.

There are 20 different naturally occurring amino acids in the human body, each of which contains a different organic side chain (also known as an R group), which is also attached to the central carbon alongside a hydrogen atom. These R groups vary widely in polarity, charge, acid-base properties and structure and will be explored in detail in VCE Units 3 & 4.

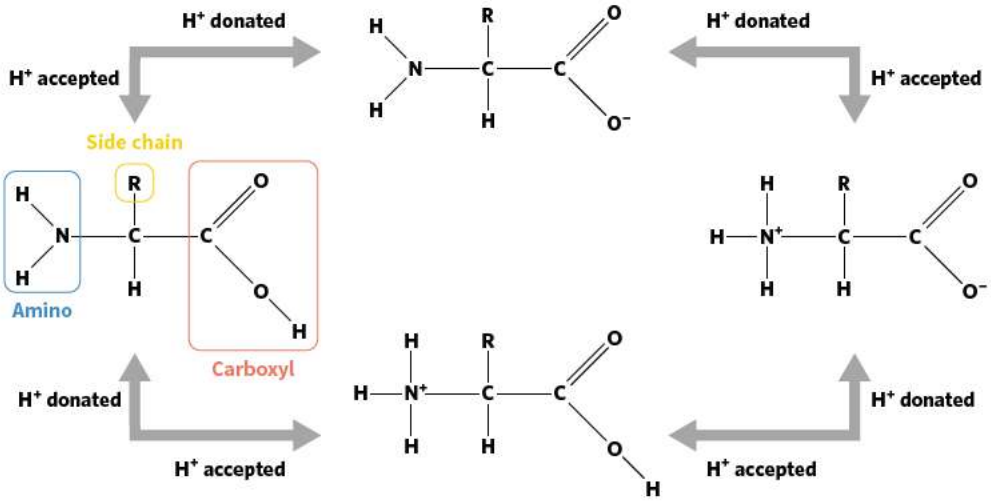


Figure 1 Amino acids can both donate and accept protons.

Amino acids react together in condensation reactions between the amino and carboxyl groups to form a **peptide** bond and a free water molecule. Peptide bonds are also known as **amide** bonds. This process is shown in figure 2.

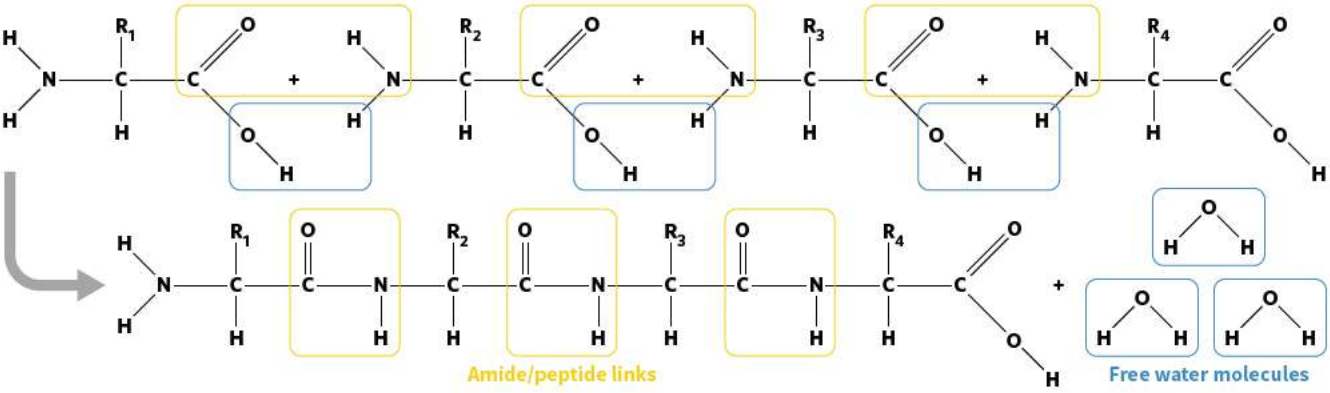


Figure 2 Amino acids combine together in condensation reactions to form proteins linked by peptide bonds.

When these amino acid chains become long enough they become known as **proteins**. These are complex structures that result from intermolecular forces such as dispersion forces, ionic interactions, permanent dipole-dipole forces and hydrogen bonding. These intermolecular forces and the structure and function of proteins will be explored in much more detail in VCE Units 3 & 4.

Enzymes

Many of the chemical reactions that occur in the human body are dependent on the presence of **enzymes**, which are specialised protein molecules that greatly increase the rate of a chemical reaction. Without these enzymes, cells would not be able to perform their vital functions within the time-frame required for life as we know it to exist. This means that enzymes are essential for the human body to function and stay alive. A number of different enzymes listed with their functions are shown in figure 3.

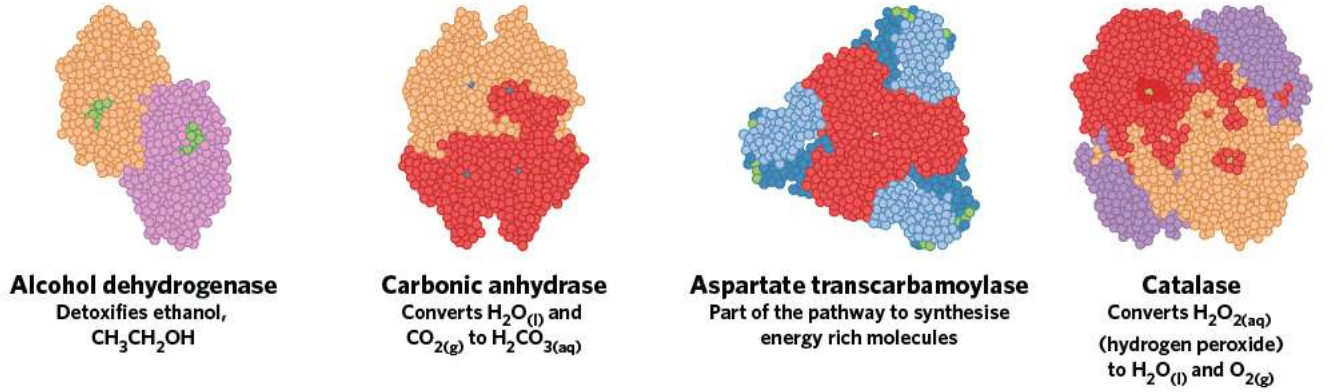


Figure 3 Enzymes have complex structures that perform a variety of important biological functions.

Since all enzymes are made up of amino acids, which have acid-base properties, they too are also affected by acid-base interactions. This means that an enzyme's ability to function is highly dependent on the pH of the surrounding environment. Because of this, enzymes require an **optimal pH** in order to function at their maximum potential. For example, enzymes in the stomach must be able to not just withstand a highly acidic pH of 2.0–3.5, but actually thrive in this environment. For most enzymes, however, they require a tightly regulated pH range of 7.35–7.45, which is the pH of human blood.

Any pH outside this range causes a large decrease in the **activity** of the enzyme, which is how quickly it can convert reactants to products. Figure 4 shows this dependence of enzyme activity on the pH of the surrounding environment.

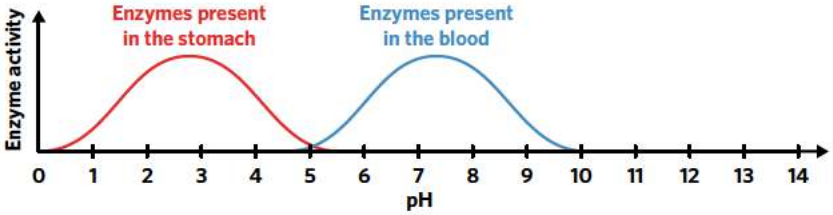


Figure 4 Enzyme activity is highly dependent on pH.

Carbonic anhydrase

One of the most important enzymes in regulating blood pH is **carbonic anhydrase**, which is present in red blood cells. Carbonic anhydrase determines the rate of the reaction to form carbonic acid, $H_2CO_{3(aq)}$, from water and $CO_{2(g)}$ dissolved in the blood. It will also increase the rate of the reaction to convert carbonic acid back into $CO_{2(g)}$ and $H_2O_{(l)}$. The carbonic acid then ionises to form acidic $H^+_{(aq)}$ ions and the bicarbonate ion, $HCO_3^-{}_{(aq)}$.

$$CO_{2(g)} + H_2O_{(l)} \overset{\text{Carbonic anhydrase}}{\rightleftharpoons} H_2CO_{3(aq)} \overset{\text{Ionisation}}{\rightleftharpoons} H^+_{(aq)} + HCO_3^-{}_{(aq)}$$

This increase in $H^+_{(aq)}$ ions results in a decrease in blood pH. The reverse reaction, which consumes carbonic acid to reform $CO_{2(g)}$ and $H_2O_{(l)}$ will result in an increase in blood pH.

Although this reaction can occur without carbonic anhydrase, the enzyme increases the rate of this reaction by a factor of 10^8 (100 000 000).

Carbon dioxide is released as a byproduct during the breakdown of sugars and fats in cells and must be removed from the body before its concentration builds up to dangerous levels. Most of the $CO_{2(g)}$ released by cells is converted to carbonic acid by carbonic anhydrase.

Since $CO_{2(g)}$ has a relatively low solubility in blood, it is primarily transported in the form of much more soluble hydrogen carbonate ions, $HCO_3^-{}_{(aq)}$. In the lungs, $CO_{2(g)}$ must be released when it is breathed out. Therefore, in the lungs it is converted from hydrogen carbonate to carbonic acid, and finally back into $CO_{2(g)}$, again by carbonic anhydrase. This cycle of $CO_{2(g)}$ transport is depicted in figure 5.

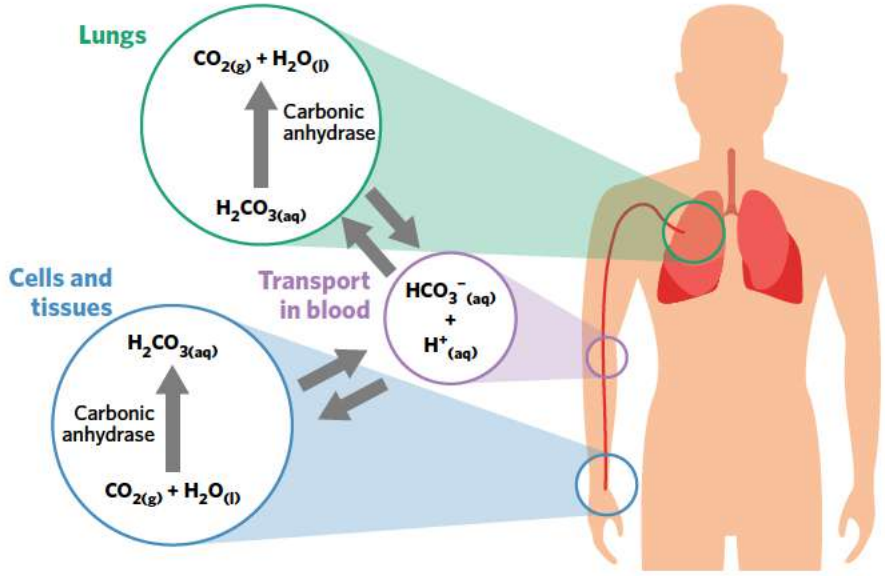


Figure 5 $CO_{2(g)}$ is transported in the blood primarily as $HCO_3^-{}_{(aq)}$. Carbonic anhydrase increases the rate of conversion between $CO_{2(g)} + H_2O_{(l)}$ and $H_2CO_{3(aq)}$.

There are a number of forms of carbonic anhydrase in the body, all of which perform different acid-base functions. In the stomach, carbonic anhydrase has a vital role in secreting acid to aid in the breakdown and digestion of food. The enzyme also allows the fluid secreted from the pancreas to remain alkaline and maintains the pH of saliva at relatively neutral. Some of the various functions of carbonic anhydrase are shown in figure 6.

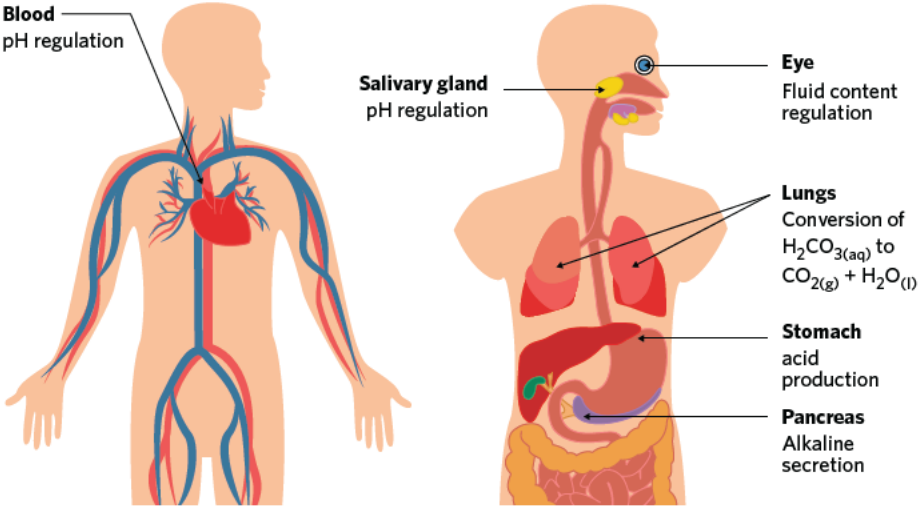


Figure 6 Carbonic anhydrase has a variety of functions in the body.

If any of these forms of carbonic anhydrase malfunction or become inactive, the body's pH regulation system will cease to function properly and the pH of the blood, saliva or stomach may reach a harmful level. This can lead to a number of different diseases and medical conditions that range from an overproduction of acid in the stomach to kidney failure and high blood acidity.

Acidosis and alkalosis

If the amount of $CO_{2(g)}$ entering the body increases or the body has difficulty removing it, the pH of the blood may decrease outside of the safe range of 7.35–7.45. One of the most common ways of this happening is through a dramatic increase or decrease in breathing rate. During normal breathing, carbon dioxide is released and oxygen is taken in.

An increase in breathing rate is known as **hyperventilation** and results in an increase in the amount of $CO_{2(g)}$ leaving the body. As this leads to a decrease in the amount of $CO_{2(g)}$ present in the body, there will be less $CO_{2(g)}$ converted to carbonic acid by carbonic anhydrase and, therefore, an increase in blood pH. When the pH of the blood becomes too alkaline or basic (above 7.45), the body enters a state known as **alkalosis**.

A decrease in breathing rate is classified as **hypoventilation** and results in a decrease in the amount of $CO_{2(g)}$ leaving the body. This will lead to an increase in the amount of $CO_{2(g)}$ that remains in the blood, also increasing the amount of $CO_{2(g)}$ converted to carbonic acid by carbonic anhydrase. This results in a decrease in blood pH as it becomes more acidic. When the pH of the blood becomes too acidic (below 7.35), the body enters **acidosis**. This relationship between breathing rate and blood pH is shown in figure 7.

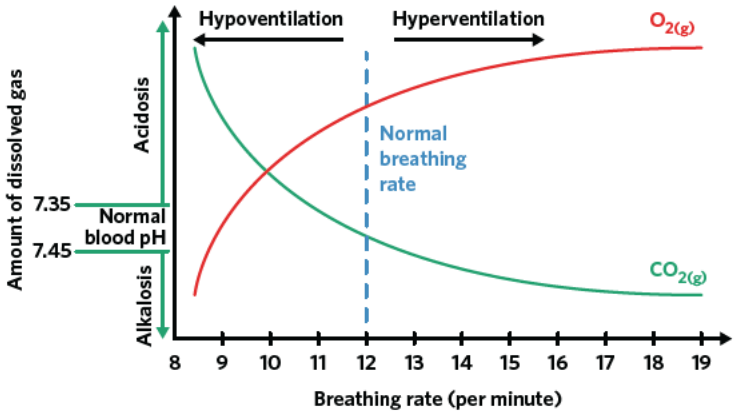


Figure 7 Deviations from the normal breathing rate will change the amount of $CO_{2(g)}$ dissolved in the blood, either increasing or decreasing blood pH.

These fluctuations in blood pH can have disastrous effects on the function of the body. The symptoms of acidosis and alkalosis due to hypoventilation and hyperventilation are summarised in table 1. If these enzymes are no longer capable of performing their vital, life-sustaining functions, the body may undergo severe damage.

Table 1 Symptoms of acidosis and alkalosis.

Acidosis	Alkalosis
Low blood pH (below 7.35)	*High blood pH (above 7.45)*
Fatigue	Nausea
Headache	Numbness
Sleepiness	Prolonged muscle spasms
Shortness of breath	Difficulty breathing
Confusion	Dizziness
Blurred vision	Coma

Theory summary

- Amino acids are amphiprotic molecules that can both donate and accept protons.
  - Amino acids are monomers, or 'building blocks', that combine together to form larger proteins linked together by peptide/amide bonds.
- Enzymes are proteins that can greatly increase the rate of chemical reactions in a biological system such as the human body.
  - Enzymes are highly sensitive to pH changes because of their acid-base properties.
- Carbonic anhydrase increases the rate of the reaction $CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$, allowing it to act as an important regulator of blood pH in the body.
  - Carbonic acid ionises according to the equation $H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)}$
  - In cells and tissues: $CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$
  - In the lungs: $H_2CO_{3(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
- Hyperventilation will lead to an increase in blood pH above 7.45, known as alkalosis.
- Hypoventilation will lead to a decrease in blood pH below 7.35, known as acidosis.

8D QUESTIONS

Theory review questions

Question 1

What are the two functional groups present in all amino acids?

A amino, carboxyl

B imino, carbonyl

Question 2

Amino acids are _____ chemical species.

A polyprotic

B amphiprotic

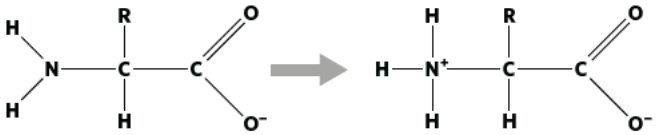
Question 3

What is the name of the functional group formed from the reaction of two amino acids?

A proline

B peptide

Question 4



In this diagram, the amino acid is acting as a/an _____.

A acid

B base

Question 5

An enzyme is a/an _____ that _____ the rate of a chemical reaction in a biological system.

A protein, increases

B amino acid, decreases

Question 6

Due to their acid-base properties, enzyme _____ is highly dependent on _____.

A concentration, temperature

B activity, pH

Question 7

What is the reaction that has its rate increased by the enzyme carbonic anhydrase?

A $H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)}$

B $CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$

Question 8

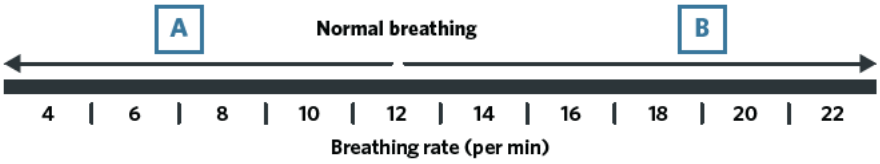
How is most carbon dioxide transported in the bloodstream?

A $CO_{2(g)}$ has a high solubility in blood and so is transported as a dissolved gas.

B $CO_{2(g)}$ has a low solubility in blood and so is transported as dissolved $HCO_3^-{}_{(aq)}$ ions.

Question 9

Which of the following correctly describes the breathing rates of A and B?



A Hypoventilation, Hyperventilation

B Hyperventilation, Hypoventilation

Question 10

Hyperventilation leads to a(n) _____ in blood CO_2, resulting in _____.

A increase, acidosis

B decrease, alkalosis

Deconstructed questions

Use the following information to answer questions 11–13.

Chronic obstructive pulmonary disease (COPD) is an umbrella term for medical conditions such as asthma, emphysema and bronchitis. Such conditions tend to increase the difficulty of breathing out (exhaling), resulting in a decreased overall breathing rate for patients affected by these conditions.

Question 11 (1 MARK)

Which of the following correctly characterises this change in breathing rate and how it will affect the amount of carbon dioxide present in the blood.

A Hyperventilation, leading to a decrease in blood $CO_{2(g)}$
B Hyperventilation, leading to an increase in blood $CO_{2(g)}$
C Hypoventilation, leading to a decrease in blood $CO_{2(g)}$
D Hypoventilation, leading to an increase in blood $CO_{2(g)}$

Question 12 (1 MARK)

Which of the following would best describe the condition of a patient suffering from bronchitis.

A The patient would have an increased risk of alkalosis.
B The patient would have a decreased risk of acidosis.
C The patient would have an increased risk of acidosis.
D The patient would have no change to their risk of acidosis or alkalosis.

Question 13 (3 MARKS)

Explain how a change in blood $CO_{2(g)}$ due to COPD would influence the pH of the blood. Justify your answer with reference to two chemical equations.

Exam-style questions

Within lesson

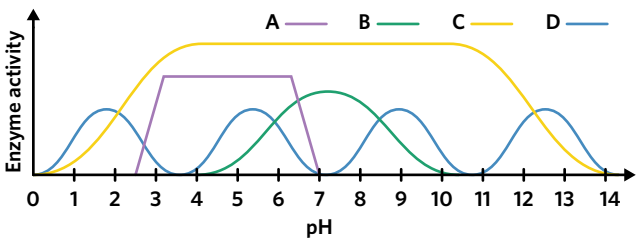
Question 14 (1 MARK)

One particular form of carbonic anhydrase found in human red blood cells is made up of 259 individual amino acids. Which of the following correctly identifies both how many peptide bonds would link carbonic anhydrase's structure together and how many free water molecules would be produced when the enzyme is synthesised?

	Peptide bonds	Water molecules
A	259	259
B	259	258
C	258	259
D	258	258

Question 15 (1 MARK)

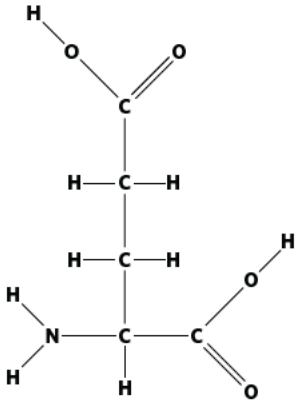
Amylase is an important enzyme in breaking down sugars such as those found in bread. It is found in the saliva, where it begins the process of chemical digestion. Which of the following curves most likely represents the relationship between amylase's activity and pH?



A Curve A
B Curve B
C Curve C
D Curve D

Question 16 (5 MARKS)

One particular amino acid, glutamic acid, has the following structure. Its side chain (R group) contains a second carboxyl group.



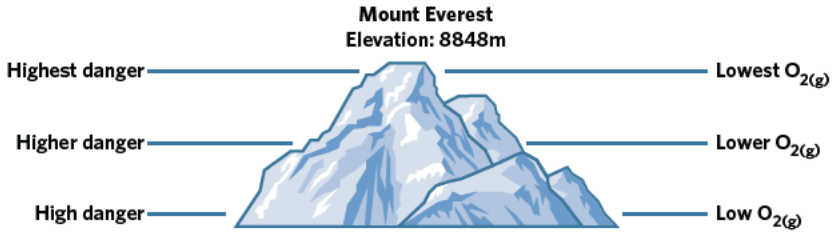
a Draw the structural formula of glutamic acid as it would appear if it has donated and accepted all possible protons. (2 MARKS)

b Would the molecule drawn in part **a** best be classified as polyprotic, monoprotic or amphiprotic? Explain. (2 MARKS)

c What would be the overall charge of the molecule drawn in part **a**? (1 MARK)

Question 17 (6 MARKS)

Charles has recently decided to scale Mount Everest, the world's tallest mountain. A high altitude environment has a decreased air pressure and a decrease in oxygen levels. However, since less oxygen is being breathed in with each breath, the body compensates for this by increasing the rate of breathing.



a What is the name given to this change in breathing rate? (1 MARK)

b Will Charles' blood $CO_{2(g)}$ content increase or decrease as he begins his ascent of Mount Everest? Explain. (2 MARKS)

c With reference to the role of carbonic anhydrase, explain how this change in blood $CO_{2(g)}$ content will impact Charles' blood pH during his expedition. (2 MARKS)

d As Charles scales his eighteenth glacier and begins to cross his fourteenth crevasse, he begins to experience severe nausea and dizziness. Are these more likely to be symptoms of acidosis or alkalosis? (1 MARK)

Multiple lessons

Question 18 (3 MARKS)

Carbonic acid is produced from $CO_{2(g)}$ and $H_2O_{(l)}$ in the body by the enzyme carbonic anhydrase. It can deprotonate according to the following equations:

$$CO_{2(g)} + H_2O_{(l)} \rightarrow \begin{cases} (1) & H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)} \\ (2) & HCO_3^-{}_{(aq)} \rightleftharpoons H^+_{(aq)} + CO_3^{2-}{}_{(aq)} \end{cases}$$

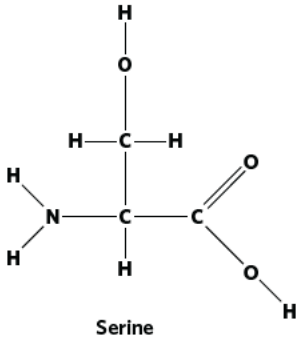
a If carbonic acid is considered a particularly weak acid, what does this suggest about the base strength of $HCO_3^-{}_{(aq)}$? (1 MARK)

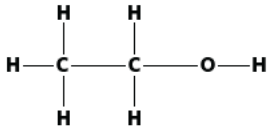
b So far in this lesson we have only considered the first deprotonation step of carbonic acid when determining blood pH. Given your answer to part **a**, would this suggest that the second deprotonation of carbonic acid occurs to any meaningful extent? What impact, if any, would this have on blood pH? (2 MARKS)

Question 19 (6 MARKS)

The structure of the amino acid serine and the alcohol ethanol are shown.

0.25 mol of serine is added to a beaker that contains 500 mL of pure ethanol, CH_3CH_2OH.


Serine


Ethanol

a Identify the intermolecular force(s) that would be present between molecules of serine and ethanol in the beaker. (1 MARK)

b Would serine dissolve in the beaker of ethanol? Explain. (2 MARKS)

c If serine is instead added to a beaker of deionised water, it is capable of both donating a proton to and accepting a proton from the water molecules.

i How would serine be classified according to its ability to donate and accept protons? (1 MARK)

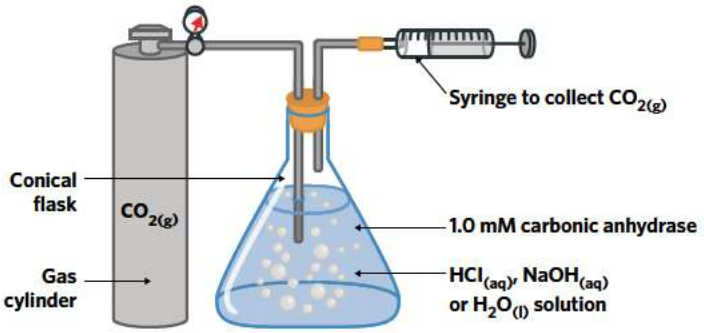
ii Write the balanced acid-base equation of serine acting as a monoprotic acid in aqueous solution. (1 MARK)

iii Identify the conjugate acid-base pair(s) that would be present in this reaction. (1 MARK)

Key science skills

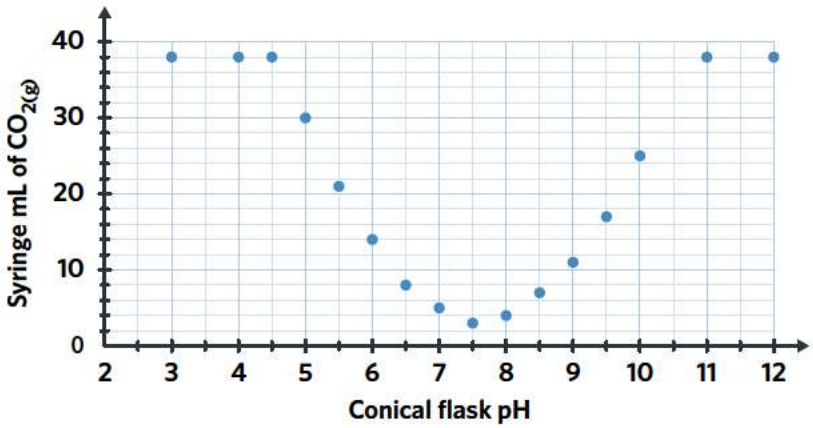
Question 20 (7 MARKS)

Fraser is interested in investigating the effect of pH on the activity of carbonic anhydrase. To do this he designs an experiment to measure how quickly carbonic anhydrase converts $CO_{2(g)}$ to carbonic acid, $H_2CO_{3(aq)}$. The experimental set up is shown.



Carbon dioxide gas is bubbled at a constant rate for 60 seconds through a solution containing 1.0 mM carbonic anhydrase and a number of different concentrations of either hydrochloric acid, sodium hydroxide or deionised water. A syringe also attached to the conical flask measures the amount of gas that escapes within the 60 seconds.

The results that Fraser collected from this experiment are shown in the following graph.



a Identify the independent variable and dependent variable in this experiment. (2 MARKS)

b How can the volume of $CO_{2(g)}$ present in the syringe be used to measure the activity of carbonic anhydrase? Explain. (3 MARKS)

c According to the results collected, at what pH is the activity of carbonic anhydrase the highest? Does this align with what you would expect for an enzyme present in the human bloodstream? (2 MARKS)

Exam-style question hints

14: Whenever a peptide bond is formed, a water molecule is released. **15:** Enzyme activity is highly dependent on a narrow pH range. **16a:** Carboxyl groups can donate protons, amino groups can accept protons. **16b:** Chemical species can be classified according to how many protons they can donate/accept. **16c:** When a neutral species donates a proton, it gains a negative charge. **17a:** Changes in breathing rate are classified as hyperventilation or hypoventilation. **17b:** The amount of $CO_{2(g)}$ in the blood depends on how much $CO_{2(g)}$ is breathed out. **17c:** Carbonic anhydrase converts $CO_{2(g)}$ to carbonic acid. **17d:** Changes in blood pH are classified as acidosis or alkalosis. **18a** Conjugate acid-base pairs have opposite strengths. **18b:** Bases accept protons, acids donate protons. **19a:** Intermolecular forces are forces of attraction that occur between two molecules. **19b:** Recall the 'like dissolves like' solubility rule. **19ci:** Serine can both donate and accept a proton. **19cii:** Acids donate protons in Brønsted-Lowry acid-base reactions. **19ciii:** Acids donate protons to form their conjugate bases. **20a** The dependent variable is affected by changes in the independent variable. **20b:** Carbonic anhydrase will convert gaseous $CO_{2(g)}$ into aqueous $H_2CO_{3(aq)}$. **20c** The activity of an enzyme is how quickly it can convert reactants to products.

EXPERIMENT

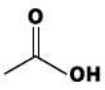
RELATIVE STRENGTH OF ORGANIC ACIDS

In this chapter, we have investigated the strength of a number of different acids. Organic acids are organic molecules with acidic properties and can vary widely in their relative strengths. The following experiment investigates the relative strengths of four different organic acids:

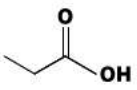
Methanoic acid



Ethanoic acid



Propanoic acid

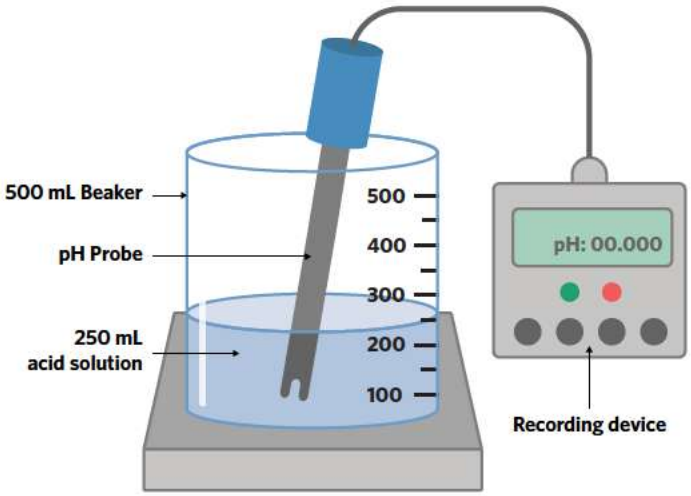


Benzoic acid



Materials

- 4 × 500 mL glass beaker
- 250 mL of 0.50 M methanoic acid, HCOOH
- 250 mL of 0.50 M ethanoic acid, CH_3COOH
- 250 mL of 0.50 M propanoic acid, CH_3CH_2COOH
- 250 mL of 0.50 M benzoic acid, C_6H_5COOH
- 1 × pH probe and recording device



Method

1. Add 250 mL of 0.50 M methanoic acid to a clean, dry 500 mL glass beaker.
2. Place the pH probe in the solution so that it is fully submerged.
3. Record the pH of the solution once the reading stabilises.
4. Repeat steps 1–3 for the CH_3COOH, CH_3CH_2COOH and C_6H_5COOH solutions.

QUESTIONS

Question 1 (1 MARK)

Identify the independent and dependent variables in this experiment.

Question 2 (2 MARKS)

Identify any controlled variables in this experiment.

Question 3 (2 MARKS)

Identify a key mistake in the experimental technique that would cause contamination of the later solutions. Suggest what could be done to prevent this.

Question 4 (5 MARKS)

The experiment was carried out by a number of different students and the following results were obtained.

	pH			
Solution	**Student 1**	**Student 2**	**Student 3**	**Average**
0.50 M HCOOH	2.033	2.027	2.022	2.027
0.50 M CH_3COOH	2.528	2.520	2.540	2.529
0.50 M CH_3CH_2COOH	2.590	2.588	2.587	2.588
0.50 M C_6H_5COOH	2.245	2.241	2.247	2.244

a Explain the reason for taking the average of multiple measurements rather than a single trial. (2 MARKS)

b List the acids in order of increasing strength. Justify your answer with reference to any relevant theory and the data collected in the experiment. (3 MARKS)

Question 5 (2 MARKS)

A student performing the experiment proposes the following hypothesis:

As the hydrocarbon chain of an unbranched carboxylic acid increases in length, the strength of the acid decreases.

Given the current design of the experiment, is this a valid hypothesis? Explain.

ANSWERS

1 Independent variable: Identity of the acid

Dependent variable: pH of the solution

2 Controlled variables:

- Volume and concentration of acid solutions
- Identity of pH probe
- Beaker in which the experiment is performed

3 [The pH probe is not cleaned between trials, allowing contamination from the previous solution in the pH reading.[1]][To prevent this, the pH probe should be cleaned with deionised water and dried before it is submerged in a new solution.[2]]

✓ ✗ I have identified the key mistake in the experimental technique.[1]

✓ ✗ I have suggested a solution to prevent this mistake.[2]

4 **a** [Repeating an experiment multiple times and averaging the results reduces the impact of any random errors in particular trials.[1]] [This increases the precision of the final values obtained.[2]]

✓ ✗ I have explained the effect of averaging results.[1]

✓ ✗ I have identified the impact of averaging results on precision.[2]

b [The stronger an acid, the greater its degree of ionisation to form $H_3O^+_{(aq)}$ ions in aqueous solution.[1]][As $H_3O^+_{(aq)}$ ion concentration increases, the pH of the solution decreases.[2]][Given that the concentrations of each monoprotic acid in the experiment are equal, a lower measured pH will correspond to a stronger acid.[3]][Therefore, in order of increasing strength: propanoic acid (2.588), ethanoic acid (2.529), benzoic acid (2.244), methanoic acid (2.027).[4]]

✓ ✗ I have explained the link between acid strength and the formation of $H_3O^+_{(aq)}$ ions.[1]

✓ ✗ I have explained the link between $H_3O^+_{(aq)}$ concentration and pH.[2]

✓ ✗ I have explained the link between acid strength and pH.[3]

✓ ✗ I have answered the question with reference to the data collected.[4]

5 [The current experimental procedure includes the use of an organic acid that is not an unbranched carboxylic acid (benzoic acid).[1]] [This renders the hypothesis invalid as the stated independent variable (length of the unbranched carboxylic acid's hydrocarbon chain) does not account for the inclusion of benzoic acid in the experimental design.[2]]

✓ ✗ I have identified the relevant part of the experimental procedure.[1]

✓ ✗ I have explained how this affects the validity of the stated hypothesis.[2]

CHAPTER 8 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Question 1 (1 MARK)

Which of the following chemical equations represents a Brønsted-Lowry acid-base reaction?

A $HPO_4^{2-}{}_{(aq)} + HBr_{(aq)} \rightarrow Br^-_{(aq)} + PO_4^{3-}{}_{(aq)}$

B $HClO_{4(aq)} + NH_{3(aq)} \rightarrow ClO_4^-{}_{(aq)} + NH_4^+{}_{(aq)}$

C $C_2H_{6(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

D $Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$

Question 2 (1 MARK)

Identify the species containing phosphorus that would be present in the reaction vessel if 12 mol of $HPO_4^{2-}{}_{(aq)}$ were added to an aqueous solution.

A $HPO_4^{2-}{}_{(aq)}$, $PO_4^{3-}{}_{(aq)}$

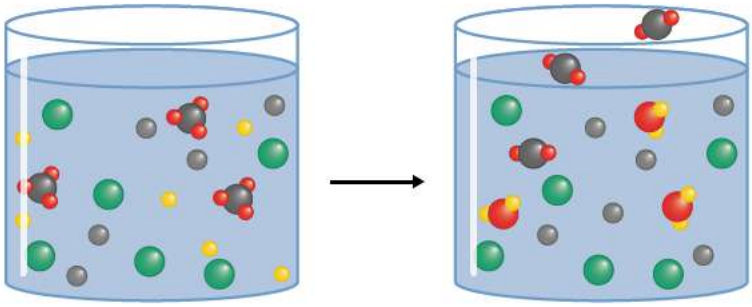
B $HPO_4^{2-}{}_{(aq)}$, $H_2PO_4^-{}_{(aq)}$, $H_3PO_{4(aq)}$

C $PO_4^{3-}{}_{(aq)}$, $HPO_4^{2-}{}_{(aq)}$, $H_2PO_4^-{}_{(aq)}$, $H_3PO_{4(aq)}$

D $PO_4^{3-}{}_{(aq)}$, $HPO_4^{2-}{}_{(aq)}$, $H_2PO_4^-{}_{(aq)}$, $H_3PO_{4(aq)}$, $H_4PO_4^+{}_{(aq)}$

Question 3 (1 MARK)

Given is a representation of a reaction occurring between an acid and a metal carbonate.



Which of the following represents the balanced ionic equation for this reaction?

A 

B 

C 

D 

Question 4 (1 MARK)

Which of the following species would be classified as both amphiprotic and polyprotic?

A $H_2PO_4^-{}_{(aq)}$

B $H_2SO_{4(aq)}$

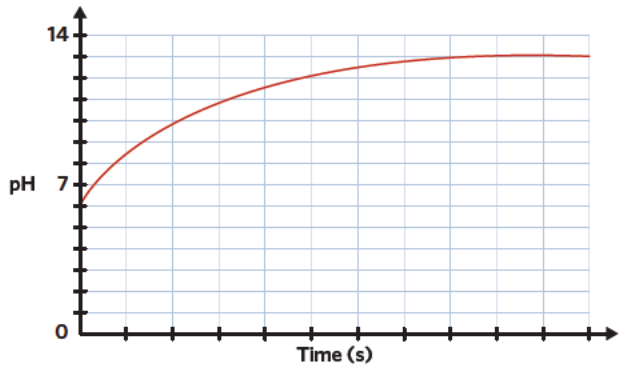
C $NO_3^-{}_{(aq)}$

D $H_3PO_{4(aq)}$

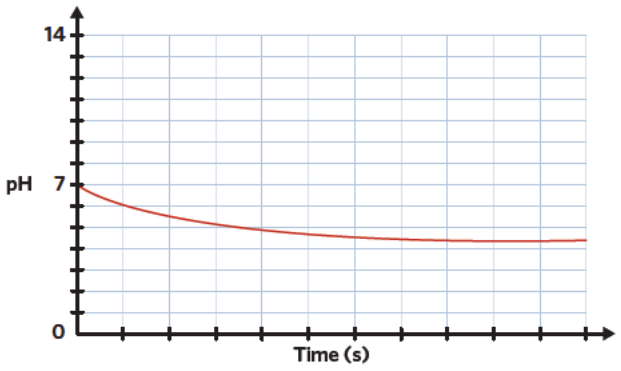
Question 5 (1 MARK)

320 g of ethanoic acid, $CH_3COOH_{(aq)}$, is added to 90.0 L of pure water. Which of the following graphs would be the most likely representation of how the solution's pH would change over time?

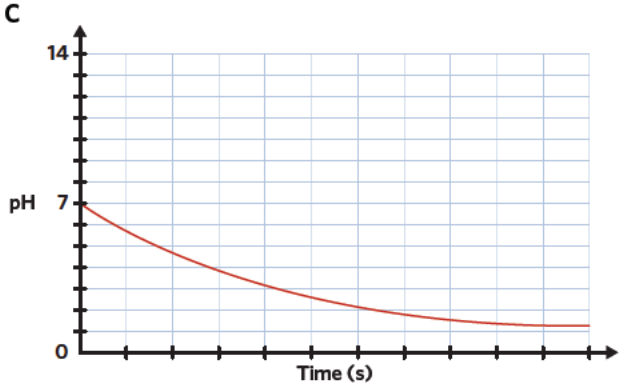
A



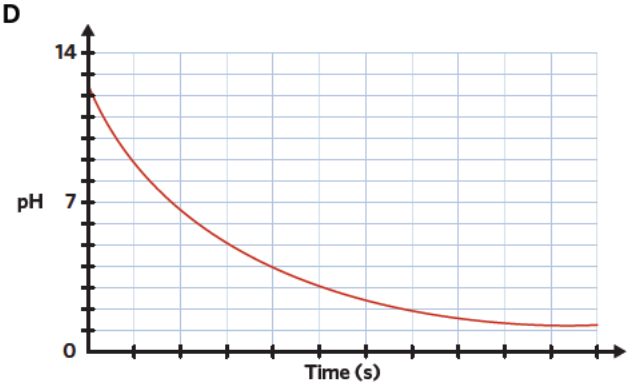
B



C



D



Question 6 (1 MARK)

$H_2SeO_{3(aq)} + H_2O_{(l)} \rightarrow HSeO_3^-{}_{(aq)} + H_3O^+{}_{(aq)}$

In this reaction $HSeO_3^-{}_{(aq)}$ is

A the conjugate base of $H_2O_{(l)}$.

B the conjugate base of $H_2SeO_{3(aq)}$.

C the conjugate acid of $H_3O^+{}_{(aq)}$.

D the conjugate acid of $H_2SeO_{3(aq)}$.

Question 7 (1 MARK)

The human stomach has an acidic pH of 3.40 due to the presence of hydrochloric acid, HCl. This destroys 99% of bacteria that are ingested, but people can still become ill and may die if they eat rotten or dangerously contaminated food. Turkey vultures rarely have this problem, as their stomach pH is slightly above zero at 0.40, which is lower than car battery acid.

Which of the following statements are **not** correct?

I Humans and turkey vultures must have different acids present in their stomachs.

II Turkey vulture stomachs are over 10 000 times more acidic than human stomachs.

III Human stomachs have a greater concentration of $OH^-{}_{(aq)}$ ions than turkey vulture stomachs.

IV Drinking a glass of water will briefly increase the pH of a human's stomach.

V Turkey vulture stomachs have a lower concentration of $H^+{}_{(aq)}$ ions than human stomachs.

A I, II, IV

B III, IV

C II, III, V

D I, II, V

Question 8 (1 MARK)

Due to a highly stressful Acids and Bases chapter review exercise, Renee has begun to hyperventilate. Which of the following best describes how this will impact her blood pH?

A The increased rate of breathing will increase the amount of $CO_{2(g)}$ in the blood, decreasing blood pH and resulting in acidosis.

B The increased rate of breathing will increase the amount of $CO_{2(g)}$ in the blood, decreasing blood pH and resulting in alkalosis.

C The increased rate of breathing will decrease the amount of $CO_{2(g)}$ in the blood, increasing blood pH and resulting in acidosis.

D The increased rate of breathing will decrease the amount of $CO_{2(g)}$ in the blood, increasing blood pH and resulting in alkalosis.

Question 9 (1 MARK)

Which of the following is the balanced overall equation for the reaction between solid magnesium and iodic acid, $HIO_{3(aq)}$?

A $HIO_{3(aq)} + Mg_{(s)} \rightarrow Mg(IO_3)_{2(aq)} + H_{2(g)}$

B $2HIO_{3(aq)} + Mg_{(s)} \rightarrow Mg(IO_3)_{2(aq)} + H_{2(g)}$

C $2H^+_{(aq)} + Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)}$

D Solid magnesium would not react with iodic acid.

Question 10 (1 MARK)

Consider the following four solutions:

I 1.9×10^{-2} M $HCl_{(aq)}$

II 17 L of water is added to a 600 mL 0.500 M $Ca(OH)_{2(aq)}$ solution

III 450 mL of 1.43 g L^{-1} $KOH_{(aq)}$

IV 24.0 L of water is added to a 70 mL 0.80 g L^{-1} $HNO_{3(aq)}$ solution

Which of the following correctly lists the four solutions in order of decreasing pH?

A II, III, IV, I

B III, II, IV, I

C II, III, I, IV

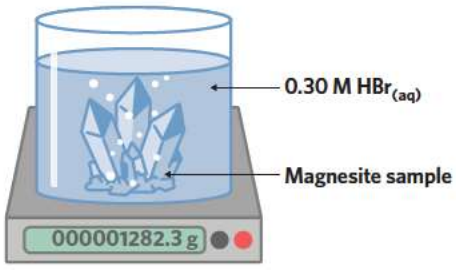
D III, II, I, IV

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (9 MARKS)

Hydrobromic acid, HBr, is an even stronger acid than hydrochloric acid. Since it is such a strong acid, it is useful in quickly extracting certain metal ores during industrial mining processes.

A chemist has recently uncovered a small deposit of magnesite in her backyard, which is a mineral with the chemical formula $MgCO_{3(s)}$. However, when she analyses her sample, she discovers that there is actually an enormous golden nugget buried within the magnesite! To remove the magnesite exterior, the chemist places the sample in a large beaker of 0.30 M hydrobromic acid.



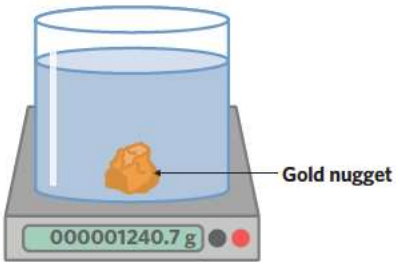
a Give the balanced overall equation for the reaction occurring in this experiment. (1 MARK)

b Give the balanced ionic equation for the reaction occurring in this experiment. (1 MARK)

c Identify the conjugate base of hydrobromic acid. If a solution of this conjugate base were added to a vessel of pure water, would the concentration of $OH^-_{(aq)}$ ions increase by any noticeable extent? Explain. (3 MARKS)

d Once all of the surrounding magnesite has been dissolved, the beaker looks as shown. Calculate the mass of magnesite that was present in the mineral sample. (2 MARKS)

e The chemist now wants to safely extract her gold nugget, but first wants to make sure that the hydrobromic acid solution is dilute enough for her to safely dispose of it. If the concentration of $H^+_{(aq)}$ ions decreased to 0.25 M after the reaction had finished, by what factor should the chemist dilute the solution so that its pH becomes equal to 5.00? (2 MARKS)



Question 12 (10 MARKS)

Evan has recently been learning how to calculate the pH of acidic and basic solutions after they have been diluted. Evan is quite skilled with the use of his calculator and can compute all his calculations rapidly and accurately. However, he often makes errors in his logical reasoning. One question Evan is given in his chemistry exam is as follows:

500 mL of 0.38 M barium hydroxide, $Ba(OH)_2$, solution is diluted by the addition of 12.0 L of water. Calculate the pH of the diluted solution if the temperature remains constant at 25°C.

Evan's working out for this calculation is shown below.

(A) Concentrated solution:

(1) $[\text{base}] = [Ba(OH)_2] = [OH^-] = 0.38\ \text{M}$

(2) $K_W = [H_3O^+][OH^-] = 10^{-14}\ \text{M}^2$

$$[H_3O^+] = \frac{10^{-14}}{0.38} = 2.63 \times 10^{-14}\ \text{M}$$

(B) Diluted solution:

(3) $c_1V_1 = c_2V_2$

$$2.63 \times 10^{-14} \times 0.500 = 12.0 \times [H_3O^+]_{\text{diluted}}$$

$$[H_3O^+]_{\text{diluted}} = \frac{2.63 \times 10^{-14} \times 0.500}{12.0} = 1.096 \times 10^{-15}\ \text{M}$$

(4) $\text{pH} = -\log_{10}[H_3O^+]$

$$\text{pH} = -\log_{10}(1.096 \times 10^{-15}) = 14.96$$

(5) The pH of the diluted solution is 14.96

Unfortunately for Evan, he has made a number of significant errors in his working out.

a Identify the mistakes made in Evan's working out and explain what you would do differently. (4 MARKS)

b Calculate the actual pH of the diluted solution. (2 MARKS)

c 3.40 L of 1.0 M $H_2SO_{4(aq)}$ is added to the diluted barium hydroxide solution. Give the balanced ionic equation for the reaction that would take place. (1 MARK)

d Would the pH of the barium hydroxide solution increase, decrease or remain the same? Explain. (2 MARKS)

e What is the name given to the reaction occurring in part **c**? (1 MARK)

Question 13 (5 MARKS)

Carbonic anhydrase is an important enzyme in regulating the pH of the blood. As part of its role, it increases the rate of one specific chemical reaction in the bloodstream that would otherwise only happen very slowly.

a Write the balanced equation for this reaction. (1 MARK)

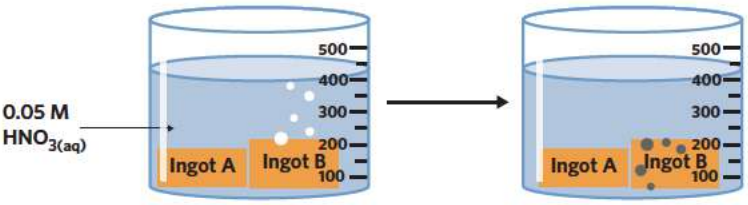
b Explain how this reaction influences the pH of the blood. Justify your answer with reference to a second chemical equation. (2 MARKS)

c Will hypoventilation increase or decrease the pH of the blood? Explain. (2 MARKS)

Question 14 (6 MARKS)

Brass is an alloy of zinc and copper, though the exact proportions of each metal vary depending on the kind of brass. Prince's metal is the name given to a version of brass that is 75% copper and 25% zinc by weight. Due to its vibrant yellow colour, Prince's metal is often used as a cheaper substitute for gold.

Unfortunately, a jeweller has mixed up an ingot of gold and an ingot of Prince's metal and, due to his poor eyesight, cannot visually tell the difference between the two. Fortunately, he remembers his Year 11 chemistry and prepares an experiment to solve his problem. The experimental set-up is shown.



a Give the balanced overall equation for the reaction(s) occurring in this experiment. (1 MARK)

b Give the balanced ionic equation for the reaction(s) occurring in this experiment. (1 MARK)

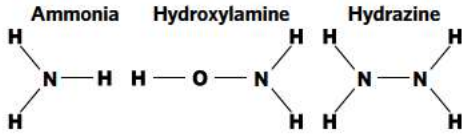
c Determine the identity of each of the two ingots. Justify your answer with reference to the experimental observations. (2 MARKS)

d When the reaction had finished, would the pH of the solution have increased, decreased or stayed the same from the beginning of the experiment? Explain. (2 MARKS)

KEY SCIENCE SKILLS QUESTIONS (10 MARKS)

Question 15 (10 MARKS)

Vignesh is in the process of designing an experiment to measure the strength of a number of different basic solutions. These bases are ammonia, NH_3, hydroxylamine, $HONH_2$, and hydrazine, N_2H_4. The structures of these molecules are shown.



Vignesh's experimental report is due tomorrow and so he has quickly assembled the following materials. However, not all of the materials will need to be used in the experiment.

Materials:

- Solutions and beakers:
  - 750 mL glass beakers (3)
  - 500 mL of 1.0 M $NH_{3(aq)}$
  - 90 mL of 1.2 M $HCl_{(aq)}$
  - 500 mL of 0.70 M $NH_{3(aq)}$
  - 250 mL of 2.0 M $N_2H_{4(aq)}$
  - 500 mL of 1.0 M $HONH_{2(aq)}$
  - 500 mL measuring cylinder
  - Deionised water
- Measurement devices:
  - Litmus paper (5 strips)
  - pH probe and recording device
  - Phenolphthalein indicator

a Write the method of a valid experiment Vignesh could conduct to determine the relative strengths of each of the bases under investigation. Ensure that as many variables are controlled for as possible. (3 MARKS)

b Explain your reasoning for using/not using the three measurement devices that Vignesh assembled. (2 MARKS)

c Explain how the data collected from the selected measurement device(s) could be used to determine the relative strengths of each of the basic solutions. (3 MARKS)

d Unfortunately, Vignesh was not able to design a valid experiment and had to search online to determine each base's strength. During his research, he discovered the following list of weak bases and their strengths.

Increasing strength ↑

butylamine	$C_4H_9NH_2$
methylamine	CH_3NH_2
ammonia	NH_3
hydrazine	N_2H_4
strychnine	$C_{21}H_{22}N_2O_2$
morphine	$C_{17}H_{19}NO_3$
hydroxylamine	$HONH_2$
pyridine	C_5H_5N
aniline	$C_6H_5NH_2$

Complete the following table using the words 'highest', 'lowest' and 'middle' to classify how the concentrations of $OH^-_{(aq)}$ and $H_3O^+_{(aq)}$ ions would appear in each solution. (2 MARKS)

	Ammonia	Hydroxylamine	Hydrazine
$[OH^-]$			
$[H_3O^+]$			

UNIT 2 AOS 1, CHAPTER 9

Redox reactions

09

9A An introduction to redox reactions

9B Redox equations

9C Metal displacement reactions

9D Redox chemistry in the body

Key knowledge

- oxidising and reducing agents, conjugate redox pairs and redox reactions including writing of balanced half and overall redox equations with states indicated
- the reactivity series of metals and metal displacement reactions including balanced redox equations with states indicated
- the causes and effects of a selected issue related to redox chemistry

Image: Mr Twister/Shutterstock.com

9A AN INTRODUCTION TO REDOX REACTIONS

In this lesson, we will be learning about the definition of redox reactions and some of the important pieces of vocabulary that are used when discussing redox reactions.

9A An introduction to redox reactions | 9B Redox equations | 9C Metal displacement reactions | 9D Redox chemistry in the body

Study design dot point

- oxidising and reducing agents, conjugate redox pairs and redox reactions including writing of balanced half and overall redox equations with states indicated

Key knowledge units

Defining redox reactions	2.1.11.1
Oxidising and reducing agents	2.1.11.2

Lesson links

This lesson builds on:

- 2A The atom

Electrons can be transferred between elements in redox reactions.

Key terms and definitions

Redox reaction chemical reaction that involves the transfer of electrons between two chemical species

Reduction chemical reaction where a chemical species gains electrons

Oxidation chemical reaction where a chemical species loses electrons

Reduced word that describes a chemical species that has undergone reduction

Oxidised word that describes a chemical species that has undergone oxidation

Half-reaction reduction or oxidation equation of a complete redox reaction

Oxidising agent (oxidant) chemical species that causes another substance to undergo oxidation and is itself reduced

Reducing agent (reductant) chemical species that causes another substance to undergo reduction and is itself oxidised

Conjugate redox pair electron acceptor (oxidising agent) and its corresponding electron donor (reducing agent)

Defining redox reactions 2.1.11.1

OVERVIEW

Redox reactions involve the transfer of electrons from one chemical species to another.

THEORY DETAILS

Redox reactions are a very common type of chemical reaction. They occur all around us. For example, the body relies on redox reactions in order to generate energy from the food we eat, and plants use redox reactions to grow via the process of photosynthesis. They also occur in non-biological systems such as the rusting of iron. All of these reactions involve the transfer of electrons from one chemical species to another, which is actually our definition of a redox reaction.

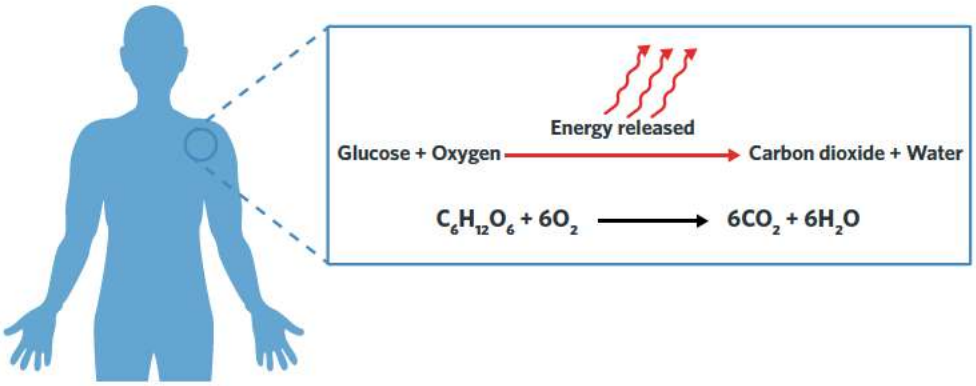


Figure 1 Important energy-producing reactions in the body are redox reactions.

The word 'redox' seems strange but it is actually a combination of two words: **reduction** and **oxidation**. As such, a redox reaction is actually the combination of a reduction and oxidation reaction occurring at the same time.

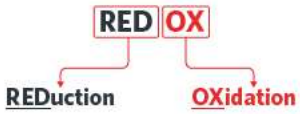


Figure 2 Origin of the word 'redox'.

Consider a reaction occurring between A and B where A is undergoing oxidation and B is undergoing reduction as shown in figure 3. In this case, compound A loses 2 electrons (oxidation), which are then given to compound B, which gains 2 electrons (reduction).

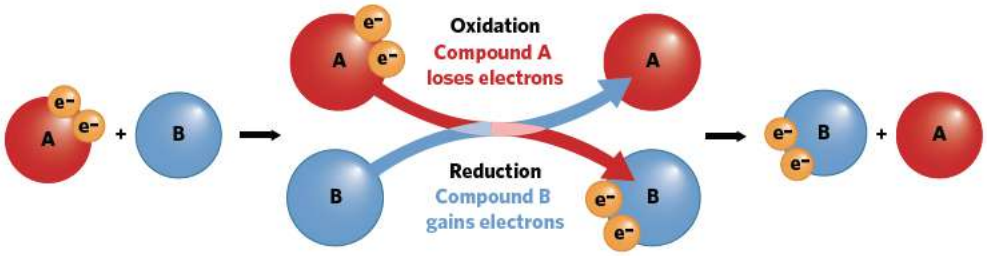


Figure 3 Model representing redox reactions.

We can show both oxidation and reduction reactions as a chemical equation that shows the movement of electrons. For example, in the following reaction $Fe^{2+}_{(aq)}$ gains two electrons, expressed as e^-, to form $Fe_{(s)}$. This is known as a reduction reaction.

$$Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$$

In this reaction, we can also say that iron ions have been **reduced**.

Conversely, in an oxidation reaction, one of the reactants loses electrons. For example, in the following oxidation reaction, $Zn_{(s)}$ loses two electrons to form $Zn^{2+}_{(aq)}$ ions.

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$$

We can also say that solid zinc has been **oxidised**.

These two reactions are known as **half-reactions** as together they make up a full redox reaction. As such, they occur simultaneously with electrons being transferred (lost and gained) between reactants to allow for redox to occur. Using the examples above, the full redox reaction can be written as:

$$Zn_{(s)} + Fe^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$$

Notice that when writing the equations for each half-reaction, electrons are shown as part of the equation. However, when writing a full equation, electrons are not shown. We will explore this concept further in the next lesson.



An easy way to remember what is happening in a redox reaction is to remember the acronym **OILRIG**.

OIL – Oxidation Is Loss of electrons

RIG – Reduction Is Gain of electrons

Another mnemonic is **LEO** the lion goes **GER**.

LEO – Loss of Electrons is Oxidation

GER – Gain of Electrons is Reduction





Worked example 1

Copper can be used for many different purposes, such as in electrical wiring. Below is a reaction involving copper.

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$

What type of half-reaction has copper taken part in?

What information is presented in the question?

A half-equation for a redox reaction.

What is the question asking us to do?

Determine whether this reaction is an oxidation or reduction half-reaction.

What strategy(ies) do we need in order to answer the question?

1. Recall the definitions of oxidation and reduction.
2. Compare the definition with the half-equation presented.

Answer

As we know, reduction reactions result in a species gaining electrons whereas oxidation results in a species losing electrons. Based on the half-equation presented, copper solid has lost electrons during the reaction to form copper ions.

As a result, this is an oxidation half-reaction.

Oxidising and reducing agents 2.11.1.2

OVERVIEW

Oxidising agents cause another chemical species to be oxidised and reducing agents cause another chemical species to be reduced.

THEORY DETAILS

As we learned in the previous section about redox reactions, they always involve two half-reactions: one oxidation and one reduction reaction both occurring at the same time. As a result, the chemical species in each half-reaction and in the whole chemical equation can be classified based on whether they cause another chemical species to be oxidised or reduced.

The word 'agent' in its most basic sense means somebody or something that performs a specific action. For example, secret agents conduct surveillance, watch suspects and other activities – they are not the ones being watched. This is an important distinction to make in order to not become confused by the vocabulary of redox reactions.

Secret agents are very active and cause actions, rather than be subjected to actions.



Figure 4 Analogy for the definition of an 'agent'.

Oxidising agents (oxidants) are chemical species which causes another chemical species to undergo oxidation. In other words, the oxidising agent causes a species to lose electrons.

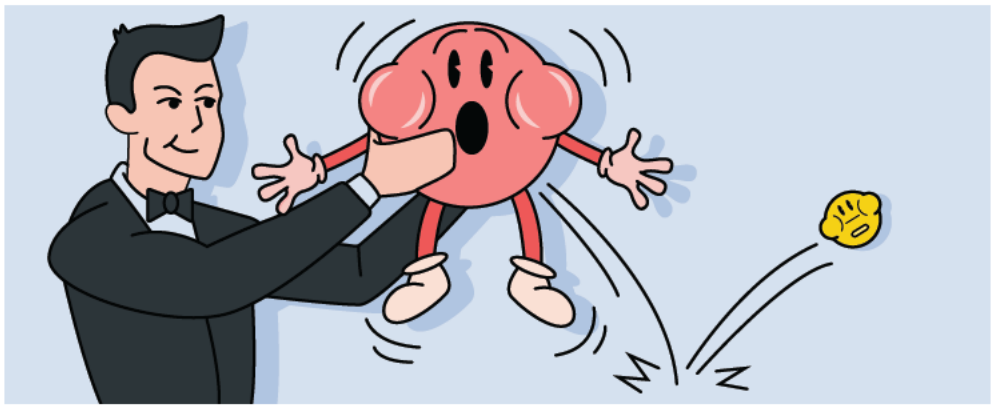
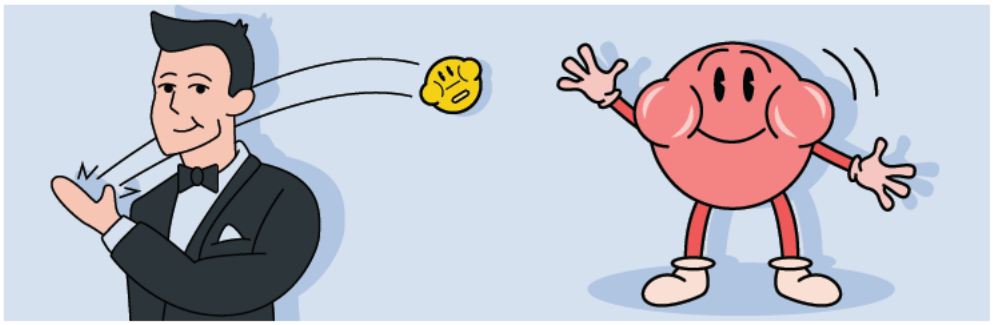


Figure 5 Illustration of how oxidising agents cause a species to lose electrons.

In contrast, **reducing agents (reductants)** are chemical species which cause another chemical species to undergo reduction. In other words, the reducing agent causes a species to gain electrons.



Oxidising agents are reduced and reducing agents are oxidised.

Figure 6 Illustration of how reducing agents cause a species to gain electrons.

A very important point to note is that oxidising agents are not oxidised and reducing agents are not reduced. Instead, oxidising agents are reduced and reducing agents are oxidised. The terminology may seem confusing but remember the analogy of an agent – they are going to perform actions rather than have actions performed upon them.

Let's take a look back at the example we had earlier of a redox reaction between zinc and iron ions.

$Zn_{(s)} + Fe^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$

We have previously shown that solid zinc undergoes oxidation to form zinc ions, and that iron ions are reduced to form solid iron. Whilst zinc solid is losing electrons, it simultaneously causes iron ions to gain electrons. This is the same for the iron ions, which gain electrons whilst simultaneously causing the solid zinc to lose electrons. From this information, we can conclude that zinc solid is a reducing agent and iron ions are the oxidising agent.

$Zn_{(s)} + Fe^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$

Reducing agent:
- Causes the reduction of another chemical species
- Is oxidised
- Loses electrons

Oxidising agent:
- Causes the oxidation of another chemical species
- Is reduced
- Gains electrons

Worked example 2

The following equation below is an example of a redox reaction:

$Mg_{(s)} + Pb^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Pb_{(s)}$

The half-reactions involved in this full redox reaction are:

$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^-$

$Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$

Identify the oxidising agent and reducing agent.

What information is presented in the question?

The complete redox reaction and the two half-equations.

What is the question asking us to do?

Determine the oxidising agent and reducing agent.

What strategy(ies) do we need in order to answer the question?

1. Determine the species being oxidised.
2. Determine the species being reduced.
3. Determine the oxidising and reducing agents.

Answer

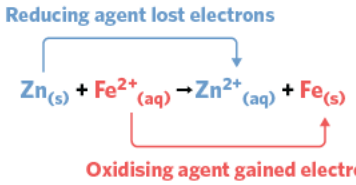
As presented in the half-reactions, we can see that solid magnesium has lost electrons and is therefore undergoing an oxidation reaction. From the second half-equation, we can see that the lead ions are gaining electrons to produce solid lead and are therefore undergoing a reduction reaction.

The solid magnesium is causing the lead ions to gain electrons (by losing electrons itself) and therefore is the reducing agent. The lead ion is causing the solid magnesium to lose electrons (by gaining electrons itself) and is therefore the oxidising agent.

Reducing agent: $Mg_{(s)}$

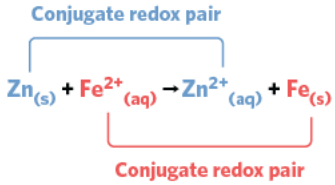
Oxidising agent: $Pb^{2+}_{(aq)}$

To summarize our example, we can see that:



Looking closely, we can see that the species in the reaction exist in pairs. In this reaction, the $Zn_{(s)}$ is oxidised to $Zn^{2+}_{(aq)}$. These two species are considered a **conjugate redox pair**. As such, $Fe^{2+}_{(aq)}$ and $Fe_{(s)}$ are also considered a conjugate redox pair.

This works similar to the acid-base conjugate pairs in lesson 8A. The reducing agent, which in this case is $Zn_{(s)}$ has a conjugate oxidising agent, $Zn^{2+}_{(aq)}$, while the oxidising agent, which in this case is $Fe^{2+}_{(aq)}$, has a conjugate reducing agent, $Fe_{(s)}$.



Worked example 3

The following equation below is an example of a redox reaction:

$Mn_{(s)} + Sn^{2+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + Sn_{(s)}$

The half-reactions involved in this full redox reaction are:

$Mn_{(s)} \rightarrow Mn^{2+}_{(aq)} + 2e^-$

$Sn^{2+}_{(aq)} + 2e^- \rightarrow Sn_{(s)}$

Identify the conjugate redox pairs.

What information is presented in the question?

The complete redox reaction and the two half-equations.

What is the question asking us to do?

Determine the conjugate redox pairs.

What strategy(ies) do we need in order to answer the question?

1. Determine the species being oxidised and therefore, the reducing agent.
2. Determine the species being reduced and therefore, the oxidising agent.
3. Determine the conjugate redox pairs.

Answer

As presented in the half-reactions, we can see that $Mn_{(s)}$ has lost electrons to produce $Mn^{2+}_{(aq)}$ and is therefore undergoing an oxidation reaction. $Mn_{(s)}$ is the reducing agent and $Mn^{2+}_{(aq)}$ is therefore the conjugate oxidising agent.
In the second half-equation, we can see that the $Pb^{2+}_{(aq)}$ ions are gaining electrons to produce $Pb_{(s)}$, and is therefore undergoing a reduction reaction. $Pb^{2+}_{(aq)}$ is therefore the oxidising agent and $Pb_{(s)}$ is its conjugate reducing agent.
Conjugate redox pairs: $Mn^{2+}_{(aq)}/Mn_{(s)}$ and $Pb^{2+}_{(aq)}/Pb_{(s)}$

When writing conjugate redox pairs, the convention is to write the oxidising agent first, followed by the reducing agent. Therefore, for our example, the conjugate pairs would be written as $Zn^{2+}_{(aq)}/Zn_{(s)}$ and $Fe^{2+}_{(aq)}/Fe_{(s)}$.

The ability to identify conjugate redox pairs is important as it helps us to determine the species undergoing oxidation and reduction in the full redox reaction.



Conjugate redox pairs are usually written with the oxidising agent first, followed by the conjugate reducing agent.

Oxidising agent/reducing agent

Theory summary

- Redox reactions are a common type of chemical reaction which involves the transfer of electrons from one chemical species to another.
- Oxidation is the loss of electrons whereas reduction is the gain of electrons.
- Oxidising agents cause the oxidation of another chemical species and are themselves reduced.
- Reducing agents cause the reduction of another chemical species and are themselves oxidised.
- Conjugate redox pairs are written as 'oxidising agent/reducing agent' and consist of an oxidising agent and its conjugate reducing agent.

9A QUESTIONS

Theory review questions

Question 1

Oxidation and reduction are entirely separate processes.

A True

B False

Question 2

Which one of the following is an oxidation half-reaction?

A $Sn^{2+}_{(aq)} + 2e^- \rightarrow Sn_{(s)}$

B $H_2O_{2(l)} \rightarrow O_{2(g)} + 2H^+_{(aq)} + 2e^-$

Question 3

For a reduction half-equation, electrons will appear on the _____ of the chemical equation because electrons are _____ during a reduction reaction.

A left-hand side, gained

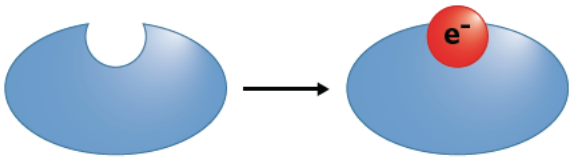
B right-hand side, lost

Question 4

What process does this diagram represent?

A Reduction

B Oxidation



Question 5

The following is a redox reaction:

$Na_{(s)} + Ag^+_{(aq)} \rightarrow Na^+_{(aq)} + Ag_{(s)}$

One of the half-equations occurring is:

$Na_{(s)} \rightarrow Na^+_{(aq)} + e^-$

Which chemical species undergoes reduction?

A $Na_{(s)}$

B $Ag^+_{(aq)}$

Question 6

The following is a redox reaction:

$Na_{(s)} + Ag^{+}_{(aq)} \rightarrow Na^{+}_{(aq)} + Ag_{(s)}$

One of the half equations occurring is:

$Na_{(s)} \rightarrow Na^{+}_{(aq)} + e^{-}$

Which chemical species is the oxidising agent?

A $Na_{(s)}$

B $Ag^{+}_{(aq)}$

Question 7

Based on the following chemical equation,

$Na_{(s)} + Ag^{+}_{(aq)} \rightarrow Na^{+}_{(aq)} + Ag_{(s)}$

which of the following options lists a conjugate redox pair?

A $Na_{(s)}/Ag^{+}_{(aq)}$

B $Na^{+}_{(aq)}/Na_{(s)}$

Question 8

Oxidising agents

A undergo oxidation.

B gain electrons.

Question 9

$Fe^{2+}_{(aq)} + Ce^{4+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + Ce^{3+}_{(aq)}$

In this reaction, the reducing agent is $Fe^{2+}_{(aq)}$. Therefore, the reductant would be

A $Ce^{4+}_{(aq)}$.

B $Fe^{2+}_{(aq)}$.

Question 10

For the following biological redox half-reaction, identify the conjugate redox pair, expressed in the correct order.

$FADH_{2(aq)} \rightarrow FAD_{(aq)} + 2H^{+}_{(aq)} + 2e^{-}$

A $FADH_{2(aq)}/FAD_{(aq)}$

B $FAD_{(aq)}/FADH_{2(aq)}$

Deconstructed questions

Use the following information to answer questions 11–13.

Aluminum is a solid metal that can react with dissolved cobalt ions in an aqueous solution. This is another example of a redox reaction.

Question 11 (1 MARK)

Redox reactions

A can either consist of an oxidation half-equation or reduction half-equation, but not both.

B involve the transfer of protons.

C involve the transfer of a negatively-charged subatomic particle.

D are the only type of chemical reaction.

Question 12 (1 MARK)

Which of the following options correctly lists the oxidising agent and reducing agent, in that order, for the chemical equation and half-equations below?

$3Co^{2+}_{(aq)} + 2Al_{(s)} \rightarrow 3Co_{(s)} + 2Al^{3+}_{(aq)}$

$Co^{2+}_{(aq)} + 2e^- \rightarrow Co_{(s)}$

$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$

A $Co^{2+}_{(aq)}$, $Al_{(s)}$

B $Al_{(s)}$, $Co^{2+}_{(aq)}$

C $Co_{(s)}$, $Al^{3+}_{(aq)}$

D $Al^{3+}_{(aq)}$, $Co_{(s)}$

Question 13 (2 MARKS)

Write out the two conjugate redox pairs present in the equation above and include a definition of the term 'conjugate redox pair'.

Exam-style questions

Within lesson

Question 14 (1 MARK)

Based on the following chemical equation and half-equations, what is the reduction half-equation in its simplest form?

$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$

$4Al_{(s)} \rightarrow 4Al^{3+}_{(s)} + 12e^-$

$3O_{2(g)} + 12e^- \rightarrow 6O^{2-}_{(s)}$

A $Al_{(s)} \rightarrow Al^{2+}_{(s)} + 2e^-$

B $Al_{(s)} \rightarrow Al^{3+}_{(s)} + 3e^-$

C $O_{2(g)} + 2e^- \rightarrow O^{2-}_{(s)}$

D $O_{2(g)} + 4e^- \rightarrow 2O^{2-}_{(s)}$

Question 15 (1 MARK)

Based on the following chemical equation and half-equations, what is the oxidation half-equation in its simplest form?

$2Au^{3+}_{(aq)} + 6I^-_{(aq)} \rightarrow 2Au_{(s)} + 3I_{2(s)}$

$2Au^{3+}_{(aq)} + 6e^- \rightarrow 2Au_{(s)}$

$6I^-_{(aq)} \rightarrow 3I_{2(s)} + 6e^-$

A $Au^{3+}_{(aq)} \rightarrow Au_{(s)} + 3e^-$

B $Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}$

C $2I^-_{(aq)} \rightarrow I_{2(s)} + 2e^-$

D $I_{2(s)} + 2e^- \rightarrow 2I^-_{(aq)}$

Question 16 (6 MARKS)

Fertilisers are mixtures of different compounds that provide the nutrients that plants need to be able to grow. Nitrogen is very important for plants because it is an atom that is found in chlorophyll, the pigment that gives plants their distinctive green colour. Therefore, a common ingredient in fertiliser is nitric acid, which can be produced by the following redox reaction:

$NH_{3(g)} + 2O_{2(g)} \rightarrow HNO_{3(g)} + H_2O_{(g)}$

a The term oxidising agent is derived from the fact that one of the first and strongest oxidising agents to be investigated was oxygen. Given this information, identify the chemical species undergoing reduction. (1 MARK)

b Identify both conjugate redox pairs present in this equation. (2 MARKS)

c Is nitric acid an oxidising or reducing agent? Explain your answer with reference to conjugate redox pairs. (2 MARKS)

The next step in the production of fertiliser is the following chemical reaction where nitric acid is neutralised by ammonia:

$HNO_{3(g)} + NH_{3(g)} \rightarrow NH_4NO_{3(s)}$

d This reaction is not a redox reaction. What does this mean in terms of the transfer of electrons between chemical species? (1 MARK)

Question 17 (8 MARKS)

A special type of redox reaction is called a disproportionation reaction. In this type of reaction, some molecules of the chemical species undergo a reduction half-equation and others undergo an oxidation half-equation. An example of this type of reaction is the following redox reaction and half-equations:

$2Cu^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + Cu_{(s)}$

$Cu^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + e^-$

$Cu^+_{(aq)} + e^- \rightarrow Cu_{(s)}$

a Explain why the same chemical species is both the oxidising agent and reducing agent. (2 MARKS)

b Write out the conjugate redox pairs for this chemical equation. (2 MARKS)

Another disproportionation reaction is the decomposition of hydrogen peroxide:

$2H_2O_{2(l)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$

c Given that $O_{2(g)}$ is considered as a conjugate oxidising agent, what is $H_2O_{2(g)}$ oxidised to? (2 MARKS)

d Hydrogen peroxide is quite unstable and so is often stored at cold temperatures and in light-free conditions. With reference to the above equation, explain why hydrogen peroxide is unstable. (2 MARKS)

Multiple lessons

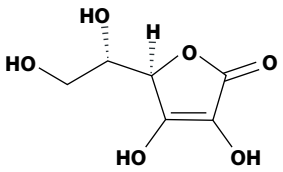
Question 18 (11 MARKS)

Some antioxidants are compounds that are used in the food industry and are studied in VCE Units 3 & 4 Chemistry. As the name suggests, antioxidants are chemical compounds that stop another chemical compound from being oxidised. They function as preservatives in foods because oxidation leads to a decline in food quality.

a If antioxidants are oxidised themselves to protect food from being oxidised, what property must they be good at? (2 MARKS)

b A student thought that since antioxidants are oxidised, they are oxidising agents. Is the student correct? Explain your answer. (2 MARKS)

Vitamin C is a naturally occurring antioxidant and its structure is shown.



c What is the name of the most common functional group in this compound? (1 MARK)

d What property of this functional group might enable vitamin C to be such a good antioxidant? (2 MARKS)

The concentration of vitamin C, $C_6H_8O_6$, can be determined by reacting it with solid iodine, $I_{2(s)}$, according to the following chemical equation and half-equations:

$I_{2(s)} + C_6H_8O_{6(aq)} \rightarrow 2I^-_{(aq)} + C_6H_6O_{6(aq)} + 2H^+_{(aq)}$

$I_{2(s)} + 2e^- \rightarrow 2I^-_{(aq)}$

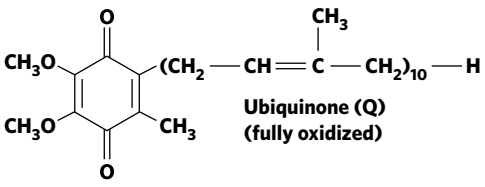
$C_6H_8O_{6(aq)} \rightarrow C_6H_6O_{6(aq)} + 2H^+_{(aq)} + 2e^-$

e Identify the oxidising agent. What is the conjugate redox pair for the oxidising agent? (2 MARKS)

f Identify the reducing agent. What is the conjugate redox pair for the reducing agent? (2 MARKS)

Question 19 (8 MARKS)

One of the ways the body generates chemical energy is through redox reactions. The process is complicated but involves important biochemical molecules. Ubiquinone (also known as coenzyme Q) is one of these molecules and, in the state pictured, it is fully oxidised.



a Can ubiquinone in this state donate or accept electrons given that it is fully oxidised? Explain your answer. (2 MARKS)

b Is ubiquinone soluble in water or a non-polar compound? Justify your answer with reference to its structure. (2 MARKS)

Free radicals are chemical compounds which contain unpaired electrons, resulting from oxidation and reduction reactions. Free radicals cause damage to the body as they are highly reactive and are suspected to be involved in many diseases.

c Suggest why a lack of coenzyme Q due to genetic causes leads to an increase in free radical concentration and can be life-threatening. (2 MARKS)

Another important biochemical molecule is Cytochrome C. It is able to carry electrons needed for the reduction of oxygen through the reduction of a Fe^{3+} ion within the molecule to Fe^{2+}.

d If Cytochrome C transports electrons to be donated to oxygen, is it functioning as a reducing or oxidising agent? (2 MARKS)

Key science skills

Question 20 (7 MARKS)

Transition metals are found approximately in the middle of the periodic table. One of their interesting features is that their varying oxidation states have different colours. This property is exploited in many chemistry experiments. Saxon wanted to determine whether a chemical species was functioning as an oxidising agent or a reducing agent when mixed with another chemical species based on these colour changes. He knows that $Ce^{4+}_{(aq)}$ is yellow, $Ce^{3+}_{(aq)}$ is colourless and $Zn^{2+}_{(aq)}$ is colourless. One result of his experiment is shown:

Chemical species mixed	Colour change observed
$Ce^{4+}_{(aq)}/Zn_{(s)}$	Yellow to colourless

a Has Saxon gathered qualitative or quantitative data? Justify your answer. (2 MARKS)

b Based on the colour change observed, what does this suggest about the reaction? (2 MARKS)

c It was known that one of the half-equations was $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$. Identify the reducing agent. (1 MARK)

d The measurement of colour changes can be subjective. Suggest a method that could be used in order to improve the reliability of this experiment so that the experiment could be reliably repeated by a different chemist. (2 MARKS)

Exam-style question hints

14: Reduction is where a chemical species gains electrons. **15:** Oxidation is where a chemical species loses electrons. **16a:** Reducing agents are themselves oxidised. **16b:** Conjugate redox pairs consist of an oxidising/reducing agent and what it is reduced/oxidised to respectively. **16c**: Oxidising agents can gain electrons. **16d:** Redox reactions are defined by the transfer of electrons from one chemical species to another. **17a:** Sometimes chemical species can be both reduced and oxidised. **17b:** Conjugate redox pairs consist of an oxidising/reducing agent and what it is reduced/oxidised to respectively. **17c** To form an oxidising agent, a chemical species is oxidised. **17d:** Substances that decompose are unstable. **18a:** Oxidation is the loss of electrons. **18b:** Oxidising agents cause another chemical to be oxidised. **18c:** Functional groups are sets of atoms which give a chemical compound unique properties. **18d**: Good antioxidants are electron-rich compounds in order for them to be able to lose electrons easily. **18e**: Oxidising agents gain electrons which are negatively charged. **18f:** Reducing agents are oxidised. **19a:** Fully oxidised compounds cannot undergo further oxidation. **19b:** Polar compounds are soluble in water (recall 'like dissolves like'). **19c:** Free radicals are formed when normal electron-carrying processes are not-functioning. **19d:** Reducing agents cause another chemical species to be reduced. **20a:** Data can either be in the form of numbers or observations. **20b:** Since the different species have different colours, the change in colour represents what is being produced. **20c:** Reducing agents cause reduction to occur **20d:** Everyone has different perceptions of colours.

9B REDOX EQUATIONS

In this lesson, we will be learning how to identify and write both half and overall redox equations.

9A An introduction to redox reactions | 9B Redox equations | 9C Metal displacement reactions | 9D Redox chemistry in the body

Study design dot point

- oxidising and reducing agents, conjugate redox pairs and redox reactions including writing of balanced half and overall redox equations with states indicated

Key knowledge units

Identifying redox equations	2.1.11.2
Writing redox half-equations	2.1.11.3
Writing overall redox equations	2.1.11.4

Key terms and definitions

Oxidation number number assigned to an atom in a substance used to determine the movement of electrons in redox reactions

Polyatomic ion ion made up of a group of atoms covalently bonded together

Permanganate ion polyatomic ion with molecular formula MnO_4^-

Dichromate ion polyatomic ion with molecular formula $Cr_2O_7^{2-}$

Lesson links

This lesson builds on:

- 9A An introduction to redox reactions

Oxidising and reducing agents form the basis of overall redox reactions.

Identifying redox reactions 2.1.11.2

OVERVIEW

Oxidation numbers are useful tools that we can use to help identify redox reactions.

THEORY DETAILS

One technique that can be used to identify redox reactions is through the use of **oxidation numbers**. Chemists created oxidation numbers as a way to show what happens to an atom during a redox reaction. More specifically, it highlights the charge that an atom may have as a result of the movement of electrons during a redox reaction.

A change in oxidation number of a chemical species indicates that a redox reaction has occurred as this means electrons have either been lost or gained. This also allows us to identify which species has undergone oxidation and which species has undergone reduction. Table 1 shows some common rules for determining the oxidation number of different forms of elements.

Table 1 Common rules to determine the oxidation number of elements.

Rule	Example
Oxidation number of a free element (an element by itself or bonded to itself) is zero	O_2, C, Cl_2
Oxidation number of a simple ion is equal to the charge of the ion	Na^+, Cl^-, O^{2-}, Al^{3+}
The sum of oxidation numbers in a neutral compound is 0	KCl, CO_2, $MgSO_4$
The sum of oxidation numbers in a polyatomic ion is equal to the charge of the ion	SO_4^{2-}, NH_4^+, MnO_4^-
The normal oxidation number for hydrogen in a compound is +1*	HCl, HF, NH_3
The normal oxidation number for oxygen in a compound is −2*	H_2O, CO, NO_3

*There are exceptions to these rules which will be introduced in VCE Chemistry Units 3 & 4.

Some of the trends in the oxidation numbers of individual elements of the periodic table are shown in figure 1.

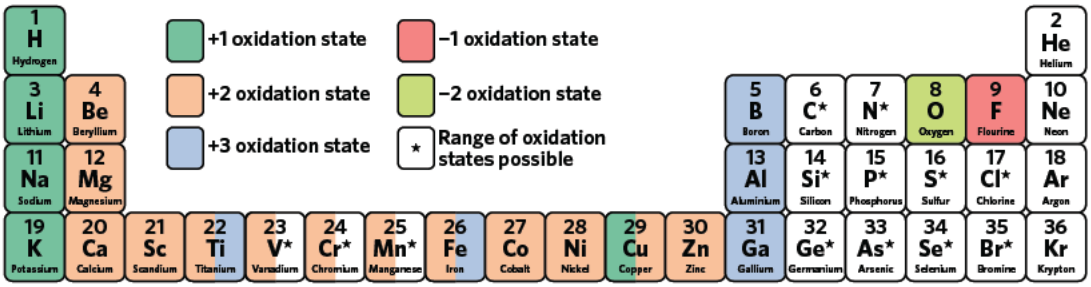


Figure 1 Periodic table showing the oxidation states of some common elements.

Let's take a look at an example of how to determine these oxidation numbers. Carbon dioxide, CO_2, is a compound that can take part in redox reactions. Based on the rules outlined in table 1, there are a couple of ideas that we need to consider:

- Carbon dioxide is a neutral compound (it has no overall charge). Therefore, the sum of oxidation numbers of the one carbon atom and the two oxygen atoms need to equal to 0.
- Each of the two oxygen atoms normally has an oxidation number of −2.

Considering we know the oxidation number of each oxygen atom, we can use this information to find the oxidation number of the carbon atom in CO_2.

CO_2 is a neutral molecule and so has an overall charge of 0

CO_2 = 0

CO_2 is composed of one C atom and two O atoms

C | O, O = 0

Each O atom will have an oxidation number of −2

? | −2, −2 = 0

The oxidation number of carbon and the two −2 oxidation numbers of oxygen will sum to 0

? + −2, −2 = 0

The two −2 oxidation numbers of oxygen contribute a total charge of −4

? + −4 = 0

Therefore, the oxidation state of the carbon atom must be +4

+4 + −4 = 0



The difference between the oxidation number and the charge of the ion is whether the 'plus' or 'minus' sign is placed before or after the number. For example the charge of the oxide ion, O^{2-}, is 2− whilst its oxidation number is −2.

Oxidation numbers are assigned for each atom of the element and not the element as a whole. Recall that the oxygen atoms in a molecule of CO_2 are not connected and so will each have their own oxidation number. In the example of CO_2, the oxidation number for oxygen is −2 not −4 as we look at each atom of the element individually, not at its total contribution to the charge. This is shown in figure 2.

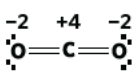


Figure 2 Oxidation numbers of carbon and oxygen in CO_2.

Worked example 1

Sulfur trioxide, SO_3, is a gas that contributes to the greenhouse effect. Determine the oxidation number of the sulfur atom in a molecule of SO_3.

What information is presented in the question?

The molecular formula of sulfur trioxide, SO_3.

What is the question asking us to do?

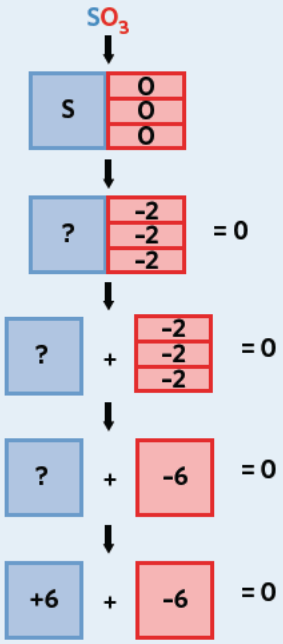
Find the oxidation number of the sulfur atom in SO_3.

What strategy(ies) do we need in order to answer the question?

1. Recall the oxidation number rule for oxygen.
2. Calculate the oxidation number of sulfur.

Answer

According to the rules in table 1, oxygen has an oxidation number of −2. However, sulfur can have a range of possible oxidation states. Sulfur trioxide is a neutral compound with an overall charge equal to 0.



Alternatively, we can calculate sulfur's oxidation number mathematically:

(oxidation number of sulfur) + 3 × (oxidation number of oxygen) = 0

$x + (3 \times -2) = 0$

$x - 6 = 0$

$x = +6$

The oxidation number of sulfur in SO_3 is +6.

While by itself, oxidation numbers do not seem to provide much information, if we pair them with our understanding of conjugate redox pairs, we can determine the species undergoing oxidation and reduction during a redox reaction.

Consider the following reaction in which carbon and oxygen react to produce carbon dioxide.

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

Based on the rules established relating to free elements, both $C_{(s)}$ and $O_{2(g)}$ will have an oxidation number of 0. However, we discovered that in a molecule of CO_2, carbon will have an oxidation number of +4 and oxygen will have an oxidation number of −2. Therefore, we can see that in this reaction there has been a change in the oxidation number of both carbon and oxygen.

$\overset{0}{C}_{(s)} + \overset{0}{O}_{2(g)} \rightarrow \overset{+4\ -2}{CO}_{2(g)}$

The changes in oxidation numbers also indicate which species has undergone oxidation and which has undergone reduction in a redox reaction. Let's first examine oxygen:

$\overset{0}{O}_{2(g)} \rightarrow \overset{-2}{CO}_{2(g)}$ Oxygen's oxidation number changes from 0 to −2.

The oxidation number of oxygen has decreased from 0 to −2. We know that redox reactions involve the movement of negatively charged electrons. Therefore, the only way for an atom to decrease in oxidation number is to gain these negatively charged electrons. This means that the oxygen atoms in $O_{2(g)}$ have gained two electrons and must have undergone reduction.

We can then examine the behaviour of carbon in this redox reaction.

$\overset{0}{C}_{(g)} \rightarrow \overset{+4}{C}O_{2(g)}$ Carbon's oxidation number changes from 0 to +4.

We can see that the oxidation number of carbon increased from 0 to +4. Once again, this change must have been due to the transfer of negatively charged electrons. For an atom to have a positive increase in oxidation number, the atom must have lost a number of negatively charged electrons. This means that the carbon atoms in $C_{(s)}$ have lost four electrons and must have undergone oxidation.



Any reaction that involves a change in the oxidation number of a species must be a redox reaction, as a change in oxidation number implies that a transfer of electrons has taken place.

As we can see from this example, oxidation numbers provide a strategy that can be used to identify:

- Whether a redox reaction has taken place, shown by a change in oxidation numbers.
- Which species has undergone oxidation, shown by an increase in oxidation number.
- Which species has undergone reduction, shown by a decrease in oxidation number.

Note that oxidation numbers and their applications will be studied in greater detail in VCE Chemistry Units 3 & 4.

Worked example 2

The permanganate ion, $MnO_4^-{}_{(aq)}$, is a common oxidising agent. Show that the unbalanced equation below is a reduction half-equation through the use of oxidation numbers.

$MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)}$

What information is presented in the question?

Unbalanced half-equation: $MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)}$

What is the question asking us to do?

Use oxidation numbers to prove that reduction is occurring.

What strategy(ies) do we need in order to answer the question?

Recall the oxidation number rules and the definition of reduction with respect to a decrease in oxidation number.

Answer

Reduction is the gain of electrons. If a chemical species gains electrons, which are negatively charged, it will experience a decrease in oxidation number.

The oxidation number of oxygen in a compound is almost always −2. The oxidation number of Mn in $MnO_4^-{}_{(aq)}$ is unknown, so we will call it x. Since the sum of the oxidation numbers in a polyatomic ion is the charge of that ion (here −1), we can generate the following equation:

MnO_4^-

$x + (4 \times -2) = -1$

$x = +7$

Therefore, the oxidation number of Mn in $MnO_4^-{}_{(aq)}$ is +7.

With respect to $Mn^{2+}{}_{(aq)}$, the oxidation number of a simple ion is equal to the charge of that ion. Therefore, the oxidation number of Mn in $Mn^{2+}{}_{(aq)}$ is +2.

Since the oxidation number of Mn has decreased from +7 to +2, this indicates that $MnO_4^-{}_{(aq)}$ has undergone reduction.

Writing redox half-equations 2.1.11.3

OVERVIEW

Redox half-equations are used to show the movement of electrons in individual oxidation and reduction reactions during a redox reaction.

THEORY DETAILS

In the previous lesson, we investigated the processes of oxidation (the loss of electrons) and reduction (the gain of electrons). These two processes occur simultaneously, with oxidation reactions coupled to reduction reactions that allow for the transfer of electrons between chemical species. A redox half-equation is a chemical equation that describes either the individual oxidation or reduction reaction that is occurring within the overall redox reaction.

As we learned in the previous section of this lesson, changes in oxidation numbers are also useful indicators of whether a redox reaction has taken place and for determining which species have undergone either oxidation or reduction.



Redox half-equations must be balanced both in the number of atoms of each species, but also in the amount of charge present on each side of the equation.

Simple redox half-equations

Consider the following oxidation reaction, where potassium solid, $K_{(s)}$, is oxidised to form potassium ions, $K^+_{(s)}$.

$$\overset{0}{K}_{(s)} \rightarrow \overset{+1}{K^+}_{(s)}$$

We know that oxidation is taking place as the oxidation number of $K_{(s)}$ increases from 0 to +1 due to the loss of a negatively charged electron. This is an increase of one. Since a loss or gain of one electron corresponds to a change in oxidation number of one, a total of one electron must have been lost for each $K_{(s)}$ atom in this oxidation reaction. This means that we can rewrite the oxidation reaction as the following half-equation, including that one electron on the products' side. Also note that the total amount of charge is balanced on both sides of the half-equation, as is the number of potassium atoms.



Remember that electrons are not molecules or atoms and so do not need to be written with a state (solid, liquid, gas or aqueous) in chemical reactions.

Half-equation: $K_{(s)} \rightarrow K^+_{(s)} + e^-$

Charge: $0 \rightarrow (+1) + (-1) = 0$

We can use a similar line of reasoning to determine how many electrons are gained during a reduction reaction, such as the reduction of chlorine gas, $Cl_{2(g)}$, to chloride ions, $Cl^-_{(aq)}$, as shown below.

$$\overset{0}{Cl}_{2(g)} \rightarrow \overset{-1}{2Cl^-}_{(aq)}$$

It is clear that reduction is taking place as the oxidation number of Cl decreases from 0 to −1, due to the gain of negatively charged electrons. Since the oxidation number of each chlorine atom decreases from 0 to −1, a total of 1 electron must have been gained per chlorine atom in the reduction reaction. As there are two chlorine atoms taking part in this reaction, the reduction reaction can, therefore, be rewritten to include 2 electrons on the reactant's side. Once again, note that the total amount of charge and the number of Cl atoms must be balanced on both sides of the half-equation.

Half-equation: $Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}$

Charge: $0 + 2 \times (-1) \rightarrow 2 \times (-1) = -2$

Worked example 3

Write the balanced redox half-equation for the conversion of $O_{2(g)}$ to $O^{2-}_{(aq)}$. Is this a reduction or oxidation reaction?

What information is presented in the question?

The molecular formulas of $O_{2(g)}$ and $O^{2-}_{(aq)}$.

What is the question asking us to do?

Write the balanced redox half-equation for the production of $O^{2-}_{(aq)}$ from $O_{2(g)}$.

What strategy(ies) do we need in order to answer the question?

1. Determine the oxidation numbers of $O_{2(g)}$ and $O^{2-}_{(aq)}$.
2. Identify whether an oxidation or reduction reaction is taking place.
3. Balance the half-equation with respect to the number of atoms and charge.

Answer

$O_{2(g)}$ is a free element and so has an oxidation number of 0.

$O^{2-}_{(aq)}$ is a simple ion, so its oxidation number is equal to its charge, −2.

$$\overset{0}{O}_{2(g)} \rightarrow \overset{-2}{O^{2-}}_{(aq)}$$

There is a decrease in oxidation number, from 0 to −2, signifying the gain of negatively charged electrons. Therefore, this is a reduction reaction.

First, we balance the number of atoms of each element:

$O_{2(g)} \rightarrow 2O^{2-}_{(aq)}$

On the reactants' side, there is a charge of 0.

On the products' side, there is a charge of $2 \times (-2) = -4$

Therefore, four electrons ($4e^-$) must be placed on the reactants' side, leading to the balanced reduction half-equation.

$O_{2(g)} + 4e^- \rightarrow 2O^{2-}_{(aq)}$

Complex redox half-equations

So far we have only considered redox half-equations that need us to balance just the number of each atom and the overall charge. We can also investigate how to balance redox-half equations that occur when **polyatomic ions** are involved. Polyatomic ions are ions that are composed of more than one atom, such as the phosphate, PO_4^{3-}, nitrate, NO_3^-, and sulfate, SO_4^{2-} ions. Just like ions that are individual atoms such as K^+ and Cl^-, polyatomic ions can also accept or donate electrons to take part in redox reactions.

Some of the most common polyatomic ions that take part in redox reactions are strong oxidising agents used in laboratory settings. These oxidising agents include the **permanganate ion**, MnO_4^-, and the **dichromate ion**, $Cr_2O_7^{2-}$. These ions are usually found as the ionic compounds potassium permanganate, $KMnO_4$, and potassium dichromate, $K_2Cr_2O_7$, which readily dissociate in aqueous solutions due to potassium's high solubility.

As these ions act as strong oxidising agents, they themselves must undergo reduction when taking part in a redox reaction. Permanganate is reduced to Mn^{2+} ions and dichromate is reduced to Cr^{3+} ions, as shown in the following two unbalanced half-equations.

$MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)}$

$Cr_2O_7^{2-}{}_{(aq)} \rightarrow Cr^{3+}{}_{(aq)}$

We can instantly see that changing just the number of manganese or chromium ions and electrons will not result in a balanced half-equation due to the presence of the multiple oxygen atoms. In fact, balancing these half-equations requires us to follow a specific sequence of balancing steps involving $H^+{}_{(aq)}$ ions and $H_2O_{(l)}$ molecules as well. To help remember the order in which these ions and molecules are added, we can use the following acronym, KOHES.

Key elements need to be balanced.
Oxygen atoms need to be balanced.
Hydrogen atoms need to be balanced.
Electrons need to be added to balance charge.
States need to be included.

Since these redox reactions are taking place in aqueous solution, we can use $H_2O_{(l)}$ to balance the number of oxygen atoms on each side of the half-equation. However, this results in an imbalance in the number of hydrogen atoms on each side of the half-equation. To balance this, we can add $H^+{}_{(aq)}$ ions to one side of the half-equation. In this lesson, we will only consider these redox reactions if they are occurring under acidic conditions, meaning that there is a ready supply of $H^+{}_{(aq)}$ ions to help us balance each half-equation.

Finally, due to the addition of positively charged $H^+{}_{(aq)}$ ions, there will be an imbalance in the amount of charge on each side of the half-equation. This is balanced through the addition of negatively charged electrons until each side has an equal amount of charge. This sequence of balancing steps is shown for the reduction of $Cr_2O_7^{2-}{}_{(aq)}$ to $Cr^{3+}{}_{(aq)}$ in figure 3.



When balancing the amount of charge in half-equations, remember that the amount of charge contributed by a chemical species in the equation depends on both the coefficient of the compound and its charge. For example, $2Cr^{3+}{}_{(aq)}$ will contribute a $2 \times 3 = +6$ total charge.

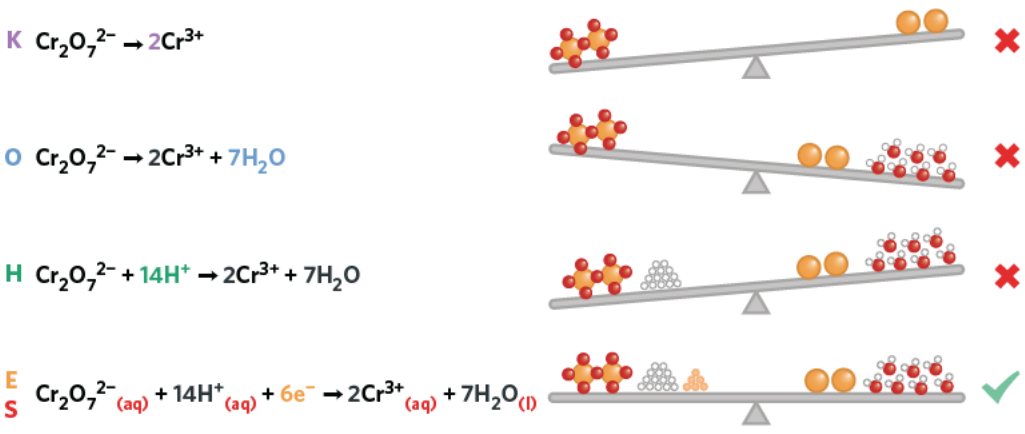


Figure 3 Balancing the reduction half-equation for $Cr_2O_7^{2-}{}_{(aq)}$ to $Cr^{3+}{}_{(aq)}$.

Worked example 4

Hydrogen sulfite, HSO_3^-, is oxidised to form the polyatomic sulfate ion, SO_4^{2-}, in an acidic aqueous solution. Write the balanced oxidation half-equation for this reaction.

What information is presented in the question?

The molecular formulas of hydrogen sulfite, HSO_3^-, and sulfate, SO_4^{2-}.

The oxidation reaction is occurring in an acidic aqueous solution.

What is the question asking us to do?

Write the balanced half-equation for the oxidation of HSO_3^- to SO_4^{2-}.

What strategy(ies) do we need in order to answer the question?

1. Balance key elements.
2. Balance oxygen atoms using H_2O.
3. Balance hydrogen atoms using H^+.
4. Balance charge using electrons.
5. Ensure states are included.

Answer

$HSO_3^- \rightarrow SO_4^{2-}$

The key element sulfur is already balanced.

$HSO_3^- + H_2O \rightarrow SO_4^{2-}$

H_2O is added to balance the four oxygen atoms on the products' side.

$HSO_3^- + H_2O \rightarrow SO_4^{2-} + 3H^+$

$3H^+$ is added to balance the three hydrogen atoms on the reactants' side.

$HSO_3^- + H_2O \rightarrow SO_4^{2-} + 3H^+ + 2e^-$

$2e^-$ is added to balance the −1 charge on the reactants' side.

$HSO_3^-{}_{(aq)} + H_2O_{(l)} \rightarrow SO_4^{2-}{}_{(aq)} + 3H^+{}_{(aq)} + 2e^-$

Appropriate states are included for all ions and covalent compounds. Note that no state is included for the electrons.

Writing overall redox equations 2.1.11.4

OVERVIEW

Overall redox equations are formed from the combination of the individual oxidation and reduction half-equations during a redox reaction.

THEORY DETAILS

Individual oxidation and reduction half-equations are useful for determining which species in a redox reaction are losing or gaining electrons. However, half-equations only represent one half of the overall redox reaction that is occurring. Since oxidation and reduction half-equations are always coupled together, an overall redox equation that combines these two half-equations can provide an overview of the reactants and products of a redox reaction.



Overall redox equations will not have electrons present on either the reactants' or products' side of the equation.

Redox reactions involve the transfer of electrons between chemical species. No additional electrons are consumed, and no additional electrons are produced during a redox reaction. Therefore, there will be the same number of electrons on both sides of an overall redox equation.

Simple overall redox equations

Consider the oxidation and reduction half-equations from the previous section of this lesson.

Oxidation: $K_{(s)} \rightarrow K^+{}_{(s)} + e^-$

Reduction: $Cl_{2(g)} + 2e^- \rightarrow 2Cl^-{}_{(s)}$

If we consider the reduction of $Cl_{2(g)}$ to $Cl^-{}_{(s)}$, a total of two electrons are consumed. However, the oxidation of $K_{(s)}$ to $K^+{}_{(s)}$ only produces one electron. This means that if we add the two half-equations together in their current form, it would appear that an additional electron is required for the reaction to proceed. This would be incorrect, as redox reactions involve instead the simultaneous transfer of electrons, with no additional electrons produced or consumed.

We have to fix this imbalance of electrons before the two half-equations are combined. We can alter the oxidation half-equation so that there will be an equal number of electrons in each half-equation by multiplying it by two as shown.

Oxidation: $2 \times (K_{(s)} \rightarrow K^+_{(s)} + e^-)$

$2K_{(s)} \rightarrow 2K^+_{(s)} + 2e^-$

Reduction: $Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(s)}$

Now that there is an equal number of electrons on each side, we can add the two half-equations together.

Overall: $2K_{(s)} + Cl_{2(g)} + 2e^- \rightarrow 2K^+_{(s)} + 2Cl^-_{(s)} + 2e^-$

Since the two electrons are present on both sides of the equation, they can be removed, leaving us with the final overall balanced redox equation.

Overall: $2K_{(s)} + Cl_{2(g)} \rightarrow 2K^+_{(s)} + 2Cl^-_{(s)}$

Complex overall redox equations

The redox half-equations of the polyatomic ions in the previous section also represent only one half of the overall redox reaction that is taking place. Consider the reduction half-equation of the permanganate ion, $MnO_4^-{}_{(aq)}$, in acidic aqueous solution.
As permanganate is a strong oxidising agent, there must be a chemical species present in the solution that it is causing to undergo oxidation. For this example, we will consider a solid piece of iron, $Fe_{(s)}$, that is oxidised to $Fe^{2+}_{(aq)}$ ions.

Oxidation $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$

Reduction: $MnO_4^-{}_{(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$

Once again, to combine these two half-equations together into the overall redox equation we need an equal number of electrons on each side. As in the previous example, the number of electrons used in reduction must equal the number of electrons released during oxidation.

In this redox reaction, two electrons are produced during the oxidation of $Fe_{(s)}$ and five electrons are consumed during the reduction of $MnO_4^-{}_{(aq)}$. Although this may seem more difficult to balance at first, all we need to do is find the lowest common multiple of five and two, which is 10. Therefore, we can multiply each half-equation by the number that will result in a total of 10 electrons being present in the half-equation.

Oxidation: $5 \times (Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-)$

$= 5Fe_{(s)} \rightarrow 5Fe^{2+}_{(aq)} + 10e^-$

Reduction: $2 \times (MnO_4^-{}_{(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)})$

$= 2MnO_4^-{}_{(aq)} + 16H^+_{(aq)} + 10e^- \rightarrow 2Mn^{2+}_{(aq)} + 8H_2O_{(l)}$

Now that there is an equal number of electrons on each side, we can add the two half-equations together.

Overall: $5Fe_{(s)} + 2MnO_4^-{}_{(aq)} + 16H^+_{(aq)} + 10e^- \rightarrow 5Fe^{2+}_{(aq)} + 2Mn^{2+}_{(aq)} + 8H_2O_{(l)} + 10e^-$

Since there are 10 electrons present on both sides of the equation, they can be removed, leaving us with the final overall balanced redox equation.

Overall: $5Fe_{(s)} + 2MnO_4^-{}_{(aq)} + 16H^+_{(aq)} \rightarrow 5Fe^{2+}_{(aq)} + 2Mn^{2+}_{(aq)} + 8H_2O_{(l)}$

Note that, unlike half-equations, overall redox equations will not have electrons present on either side of the equation.



Electrons, as well as any other chemical species (such as $H^+_{(aq)}$ or $H_2O_{(l)}$) that appears on both sides of a chemical equation, can be removed during the balancing process.



When multiplying an equation by some factor, remember that it is not only the number of electrons that is multiplied. All other chemical species will also be multiplied by the same number.

Worked example 5

A piece of solid cadmium is added to a solution of acidified potassium dichromate, resulting in the following two half-reactions occurring.

(1) $Cr_2O_7^{2-}{}_{(aq)} + 14H^+{}_{(aq)} + 6e^- \rightarrow 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$

(2) $Cd_{(s)} \rightarrow Cd^{2+}{}_{(aq)} + 2e^-$

What is the balanced overall redox equation for the reaction taking place?

What information is presented in the question?

The balanced half-equations for oxidation and reduction.

What is the question asking us to do?

Determine the balanced overall redox equation for this reaction.

What strategy(ies) do we need in order to answer the question?

1. Ensure that there is an equal number of electrons in each half-equation.
2. Add the two reactions together.
3. Remove any chemical species that are present on both sides of the overall redox equation.

Answer

The reduction half-equation of $Cr_2O_7^{2-}{}_{(aq)}$ shows that $6e^-$ are consumed. The oxidation half-equation for $Cd_{(s)}$ shows that only $2e^-$ are being produced. We require an equal number of electrons to be consumed and produced.

Therefore, half-equation (2) must be multiplied by a factor of three.

$3 \times (Cd_{(s)} \rightarrow Cd^{2+}{}_{(aq)} + 2e^-)$

$= 3Cd_{(s)} \rightarrow 3Cd^{2+}{}_{(aq)} + 6e^-$

We can now add this half-equation to half-equation (1).

$Cr_2O_7^{2-}{}_{(aq)} + 14H^+{}_{(aq)} + 3Cd_{(s)} + 6e^- \rightarrow 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)} + 3Cd^{2+}{}_{(aq)} + 6e^-$

As there are six electrons present on both sides of the equation, they must be removed, leaving the final balanced overall redox equation.

$Cr_2O_7^{2-}{}_{(aq)} + 14H^+{}_{(aq)} + 3Cd_{(s)} \rightarrow 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)} + 3Cd^{2+}{}_{(aq)}$

Theory summary

- Oxidation numbers are values assigned to the atoms in molecular compounds to track the movement of electrons between chemical species in redox reactions.
- Changes in oxidation numbers can be used to determine whether a redox reaction has taken place.
  - An increase in oxidation number means that species has undergone oxidation.
  - A decrease in oxidation number means that species has undergone reduction.
- Redox half-equations describe the individual oxidation and reduction reactions occurring within a redox reaction.
  - Redox half-equations will show the movement (loss or gain) of electrons.
- Redox half-equations and overall redox equations must be balanced with respect to both the number of each atom and the amount of charge present on each side of the equation.
  - The balancing of redox half-equations in acidic solution follows a specific sequence of steps described by the acronym KOHES.
  - The permanganate ion, MnO_4^-, and the dichromate ion, $Cr_2O_7^{2-}$, are polyatomic ions that are strong oxidising agents.
- Overall redox equations are formed from the combination of the oxidation and reduction half-equations.
  - Overall redox equations will not explicitly show the movement (loss or gain) of electrons.

9B QUESTIONS

Theory review questions

Question 1

During a redox reaction, there will always be a change in a chemical species' oxidation number.

A True

B False

Question 2

The sum of the oxidation numbers in the compound K_2SO_4 will be equal to _____.

A 0

B −2

Question 3

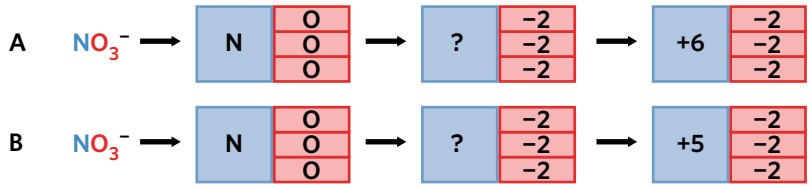
In most compounds, hydrogen will have an oxidation number of _____ and oxygen will have an oxidation number of _____.

A −1, +2

B +1, −2

Question 4

Which of the following best represents the oxidation numbers of nitrate, NO_3^-?



Question 5

Is the oxidation number of the phosphate atom in the phosphate ion, PO_4^{3-}, positive or negative?

A Positive

B Negative

Question 6

An increase in the oxidation number of a chemical species means that it has undergone _____.

A oxidation

B reduction

Question 7

Redox half-equations only need to be balanced with respect to the number of each element, but not necessarily with respect to charge.

A True

B False

Question 8

Which of the following represents an overall redox equation?

A $O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$

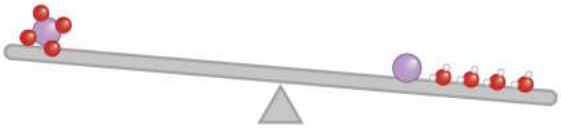
B $2H_2O_{(l)} \rightarrow O_{2(g)} + 2H_{2(g)}$

Question 9

Consider the following redox half-equation.

$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

What would be the next step in balancing this redox half-equation?



A Adding 8 ○ to the left-hand side of the half-equation

B Adding 2 ● to the right-hand side of the half-equation

Question 10

Which of the following represents a balanced redox half-equation?

A $SO_4^{2-}{}_{(aq)} + 2H^+_{(aq)} \rightarrow SO_{2(g)} + H_2O_{(l)}$

B $SO_4^{2-}{}_{(aq)} + 4H^+_{(aq)} + 2e^- \rightarrow SO_{2(g)} + 2H_2O_{(l)}$

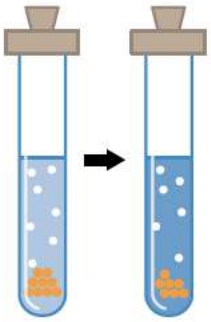
Deconstructed questions

Use the following information to answer questions 11–13.

If solid copper is added to a solution of nitric acid, $HNO_{3(aq)}$, the colourless gas nitric oxide, $NO_{(g)}$, can be formed. The solution will also transition from colourless to vibrant blue due to the formation of $Cu^{2+}_{(aq)}$ ions.

The unbalanced overall redox equation for this reaction is given.

$Cu_{(s)} + NO_3^-{}_{(aq)} \rightarrow NO_{(g)} + Cu^{2+}_{(aq)}$



Question 11 (1 MARK)

Which of the following correctly identifies the oxidation numbers of all chemical species present in this reaction?

A $\overset{0}{Cu}_{(s)} + \overset{+6\,-2}{NO_3^-}{}_{(aq)} \rightarrow \overset{+2\,-2}{NO}_{(g)} + \overset{+2}{Cu^{2+}}_{(aq)}$

B $\overset{0}{Cu}_{(s)} + \overset{+5\,-2}{NO_3^-}{}_{(aq)} \rightarrow \overset{+2\,-2}{NO}_{(g)} + \overset{+2}{Cu^{2+}}_{(aq)}$

C $\overset{0}{Cu}_{(s)} + \overset{+2\,-1}{NO_3^-}{}_{(aq)} \rightarrow \overset{+1\,-1}{NO}_{(g)} + \overset{+2}{Cu^{2+}}_{(aq)}$

D $\overset{-2}{Cu}_{(s)} + \overset{+5\,-2}{NO_3^-}{}_{(aq)} \rightarrow \overset{+2\,-2}{NO}_{(g)} + \overset{+2}{Cu^{2+}}_{(aq)}$

Question 12 (1 MARK)

Which of the following correctly classifies the species undergoing oxidation and reduction in this redox reaction?

	Oxidation	Reduction
A	$Cu_{(s)}$	$NO_3^-{}_{(aq)}$
B	$Cu^{2+}_{(aq)}$	$NO_{(g)}$
C	$NO_3^-{}_{(aq)}$	$Cu_{(s)}$
D	$NO_{(g)}$	$Cu^{2+}_{(aq)}$

Question 13 (2 MARKS)

By first writing the balanced oxidation and reduction half-equations for this reaction, give the balanced overall redox equation for the reaction between nitric acid and solid copper.

Exam-style questions

Within lesson

Question 14 (1 MARK)

$IO_3^-{}_{(aq)} \rightarrow I^-{}_{(aq)} + 3H_2O_{(l)}$

Which of the following best describes the given unbalanced reaction?

A This is an oxidation reaction as the oxidation number of oxygen decreases from −2 to −3.

B This is a reduction reaction as the oxidation number of iodine decreases from to +7 to +1.

C This is a reduction reaction as the oxidation number of iodine decreases from +5 to −1.

D This is not a redox reaction.

Question 15 (1 MARK)

Which of the following is **not** a redox reaction?

A $2HClO_{4(aq)} + Mg_{(s)} \rightarrow Mg(ClO_4)_{2(aq)} + H_{2(g)}$

B $H_{2(g)} + F_{2(g)} \rightarrow 2HF_{(g)}$

C $2C_5H_{10(g)} + 15O_{2(g)} \rightarrow 10CO_{2(g)} + 10H_2O_{(l)}$

D $CdSO_{4(aq)} + K_2S_{(aq)} \rightarrow CdS_{(s)} + K_2SO_{4(aq)}$

Question 16 (36 MARKS)

Redox half-equations and overall redox-equations can be used to demonstrate the transfer of electrons between chemical species in a redox reaction.

a Balance each of the following half-equations. Identify whether the reaction occurring is an oxidation or reduction process.

i $Br_{2(l)} \rightarrow Br^-{}_{(aq)}$ (2 MARKS)

ii $Cr^{3+}{}_{(aq)} \rightarrow Cr_{(s)}$ (2 MARKS)

iii $Ag_{(s)} \rightarrow Ag^+{}_{(aq)}$ (2 MARKS)

iv $I^-{}_{(aq)} \rightarrow I_{2(s)}$ (2 MARKS)

v $Sn^{4+}{}_{(aq)} \rightarrow Sn^{2+}{}_{(aq)}$ (2 MARKS)

vi $2H_2O_{(l)} \rightarrow H_2O_{2(aq)} + 2H^+{}_{(aq)}$ (2 MARKS)

b Balance each of the following half-equations as they would occur in an acidic solution. Identify whether the half-reaction occurring is an oxidation or reduction process.

i $PO_3^{3-}{}_{(aq)} \rightarrow PO_4^{3-}{}_{(aq)}$ (2 MARKS)

ii $Mn_2O_{3(aq)} \rightarrow MnO_{2(s)}$ (2 MARKS)

iii $NO_3^-{}_{(aq)} \rightarrow N_2O_{4(g)}$ (2 MARKS)

iv $Pb^{2+}{}_{(aq)} \rightarrow PbO_{2(s)}$ (2 MARKS)

v $ReO_4^-{}_{(aq)} \rightarrow Re_{(s)}$ (2 MARKS)

vi $NH_{3(aq)} \rightarrow NO_{2(g)}$ (2 MARKS)

c Write the balanced overall redox equations for the reactions that would occur between each of the following combinations. If this redox reaction would not occur, state 'No reaction'. Remember that any chemical species or electrons present on both sides of an overall redox equation must be removed.

i a.i. & a.iv. (1 MARK)

ii a.ii. & a.iii. (1 MARK)

iii a.v. & b.iii. (1 MARK)

iv a.vi. & b.iii. (1 MARK)

v b.ii. & b.v. (1 MARK)

vi a.v. & b.iv. (1 MARK)

d Balance each of the following redox equations by first identifying and balancing the individual half-equations, then combining them to form the overall redox equation.

i $Cr_2O_7^{2-}{}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + Cr^{3+}{}_{(aq)}$ (2 MARKS)

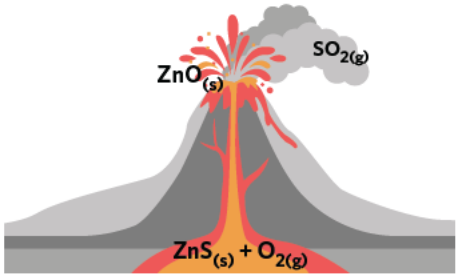
ii $I_{2(s)} + PbO_{2(s)} \rightarrow Pb^{2+}{}_{(aq)} + IO_3^-{}_{(aq)}$ (2 MARKS)

iii $As_{(s)} \rightarrow AsH_{3(s)} + H_2AsO_4^-{}_{(aq)}$ (2 MARKS)

Question 17 (5 MARKS)

Sulfur dioxide, $SO_{2(g)}$, is a toxic gas released naturally during volcanic eruptions. One of the main reactions responsible for this release of thousands of tonnes of $SO_{2(g)}$ is the roasting of sphalerite ores, also known as zinc sulfide, $ZnS_{(s)}$.
The balanced chemical reaction for this process is shown.

$2ZnS_{(s)} + 3O_{2(g)} \rightarrow 2ZnO_{(s)} + 2SO_{2(g)}$



a In the conversion of $ZnS_{(s)}$ to $SO_{2(g)}$, does sulfur undergo oxidation or reduction? Explain. (2 MARKS)

b Write the balanced half-equation for the conversion of $S^{2-}{}_{(s)}$ to $SO_{2(g)}$ as it would occur in acidic aqueous solution. (1 MARK)

c Write the balanced half-equation for the conversion of $O_{2(g)}$ to $O^{2-}{}_{(s)}$. (1 MARK)

d For every mole of $ZnS_{(s)}$ consumed, how many moles of $ZnO_{(s)}$ will be produced? (1 MARK)

Multiple lessons

Question 18 (10 MARKS)

Identify the oxidising and reducing agent in the redox reactions given in parts **a** and **b**. Justify your answers with reference to oxidation numbers.

a $2FeCl_{3(s)} + H_2S_{(g)} \rightarrow 2FeCl_{2(s)} + S_{(s)} + 2HCl_{(g)}$ (2 MARKS)

b $3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$ (2 MARKS)

c Write the balanced oxidation and reduction half-equations for these two redox reactions. For ionic compounds, include only the ion undergoing a change in oxidation number (e.g. only the $Fe^{3+}{}_{(s)}$ or $Cl^{-}{}_{(s)}$ in $FeCl_{3(s)}$). (4 MARKS)

d Identify the conjugate redox pairs for the reactions in parts **a** and **b**. (2 MARKS)

Question 19 (12 MARKS)

Chlorate, $ClO_3^-{}_{(aq)}$, is an incredibly strong oxidising agent that was once commonly used in pyrotechnic (e.g. firework) displays due to its explosive reactivity with organic compounds. When taking part in a redox reaction in an acidic solution, the chlorate ion will be reduced to form chlorine gas, $Cl_{2(g)}$.

a Write the balanced half-equation for the reduction of $ClO_3^-{}_{(aq)}$ to $Cl_{2(g)}$. Identify the oxidation numbers for all atoms present in the half-equation. (4 MARKS)

b Chlorate is capable of oxidising nitric oxide, $NO_{(g)}$, into nitrate ions, $NO_3^-{}_{(aq)}$. Give the balanced oxidation half-equation for this reaction. Identify the oxidation numbers for all atoms present in the half-equation. (4 MARKS)

c Using these two half-equations, write the balanced overall redox equation for the reaction between chlorate and nitric oxide. (2 MARKS)

d Would you expect the pH of the solution in which this reaction was occurring to increase, decrease or remain constant as a result of this reaction? Explain. (2 MARKS)

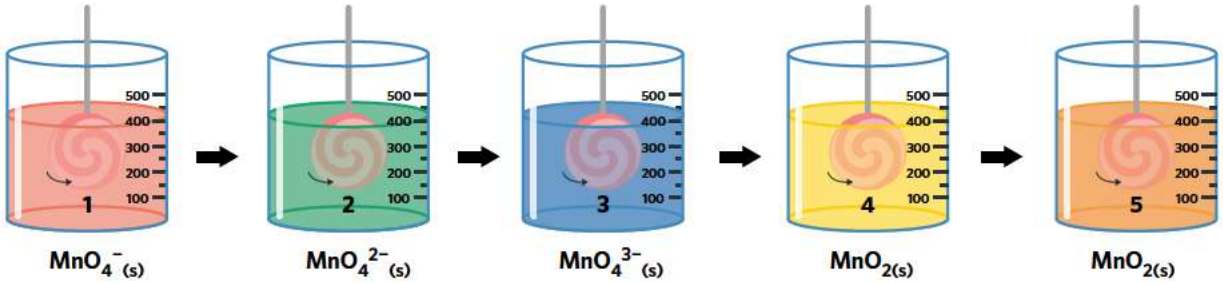
Key science skills

Question 20 (6 MARKS)

Manganese is a transition metal capable of existing in a wide range of oxidation states, which is the name given to the state of an atom at a particular oxidation number. For example, if the manganese atom in a compound has an oxidation number of +4, the atom is said to be in the +4 oxidation state. Depending on manganese's oxidation state in a compound, it can form a number of differently coloured solutions.

Pallas is a chemistry student interested in investigating these colour changes through the use of a potassium permanganate, $KMnO_{4(aq)}$, solution, which he knows is a strong oxidising agent. He also knows that the sugars present in confectionery such as lollipops can act as reducing agents.

Pallas designs an experiment in which he stirs a lollipop in an alkaline solution of $KMnO_{4(aq)}$ and records any colour changes he observes throughout the experiment. A diagram of his experiment and results is shown.



a Identify the dependent variable in this experiment. (1 MARK)

b Complete Pallas's table of results. (2 MARKS)

	1	2	3	4	5
Colour	Pink	Green	Blue	Yellow	Orange
Main species present	$MnO_4^-{}_{(aq)}$	$MnO_4^{2-}{}_{(aq)}$	$MnO_4^{3-}{}_{(aq)}$	$MnO_{2(s)}$	$MnO_{2(s)}$
Mn oxidation state					

c Give the balanced redox half-equation for the reaction responsible for the solution's colour transition from green to blue. (1 MARK)

d Suggest a reason for why there is a second colour observed in the beaker at point 5, even though there was the same chemical species present. (2 MARKS)

Exam-style question hints

14: Reduction reactions result in a decrease in oxidation number. **15:** Redox reactions involve changes in oxidation numbers. **16:** Half-equations and overall equations must be balanced according to a specific sequence of steps. **17a:** Changes in oxidation numbers help determine whether electrons have been lost or gained. **17b:** Half-equations must be balanced according to a specific sequence of steps. **17c:** Half-equations must be balanced according to a specific sequence of steps. **17d:** A reaction's stoichiometric ratio describes the ratios of reactants and products. **18a:** Oxidising agents themselves undergo reduction. **18b:** Oxidising agents themselves undergo reduction. **18c:** Half-equations show the loss or gain of electrons. **18d:** Conjugate redox pairs consist of the oxidising agent and its conjugate reducing agent. **19a:** Half-equations must be balanced according to a specific sequence of steps. **19b:** Half-equations must be balanced according to a specific sequence of steps. **19c** Overall redox reactions are formed from their individual half-equations. **19d:** pH is proportional to the number of H^+ ions present in a solution. **20a:** The dependent variable is measured in an experiment. **20b:** Oxidation state is equal to oxidation number. **20c:** Different Mn compounds are observed through colour changes. **20d:** The amount of a compound in a solution can affect its colour.

9C METAL DISPLACEMENT REACTIONS

In this lesson, we will learn about the reactivity series of metals and how it can be used to identify and explain a specific class of redox reactions known as metal displacement reactions.

9A An introduction to redox reactions | 9B Redox equations | 9C Metal displacement reactions | 9D Redox chemistry in the body

Study design dot point

- the reactivity series of metals and metal displacement reactions including balanced redox equations with states indicated

Key knowledge units

Reactivity series of metals	2.1.12.1
Metal displacement reactions	2.1.12.2

Key terms and definitions

Reactivity series of metals organised depiction of metals and their cations, ranked according to their strength as reducing and oxidising agents

Metal displacement reaction redox reaction in which a more reactive metal displaces a less reactive metal's cations from solution

Spontaneous redox reaction redox reaction that occurs without the addition of any extra heat or energy

Lesson links

This lesson builds on:

- 3C Reactivity of metals

The varying reactivity of different metals determines whether metals and their cations will react together in a metal displacement reaction.

Reactivity series of metals 2.1.12.1

OVERVIEW

Metals and their cations can be organised according to their relative strengths as reducing and oxidising agents which is the basis for the reactivity series of metals.

THEORY DETAILS

As we learned in chapter 3, metals can be classified according to their reactivity with oxygen, cold water, steam and dilute acids. All of these reactions are examples of redox reactions. In this lesson, we will investigate the oxidation and reduction processes enabling these reactions to occur and how metals can be organised according to their reactivity.

When a metal in its solid form reacts with oxygen, water or an acid, it acts as a reducing agent, losing electrons as it forms its corresponding metal cation. We can then rank these metals from the weakest reducing agent to the strongest reducing agent. This is known as the **reactivity series of metals**. The reactivity series of a number of common metals is given in figure 1, showing the reduction half-equations for each metal's cation

Metal cations are shown on the left-hand side of the reactivity series, with the weakest oxidising agent (Li^+) in the bottom-left position. The strength of the metal cation as an oxidising agent increases from bottom to top, with the strongest oxidising agent (Pt^{2+}) present in the top-left position.

Metals in their solid form are shown on the right-hand side of the reactivity series, with the weakest reducing agent (Pt) in the top-right position. The strength of the metal as a reducing agent will increase from top to bottom, with the strongest reducing agent (Li) present in the bottom-right position.

Remember that the stronger a reducing agent, the weaker its conjugate oxidising agent will be, as shown in figure 1. The inverse also applies – the stronger an oxidising agent, the weaker its conjugate reducing agent.

The more reactive a metal is, the more readily it undergoes oxidation, reacting vigorously and readily with oxygen, water and dilute acids. A metal like gold is valued highly for both its colour and because it is very unreactive. It will not react and become degraded by oxygen or water.

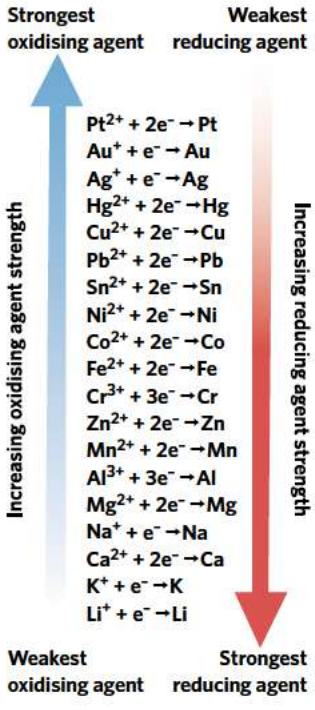


Figure 1 Reactivity series of metals.

A highly reactive metal such as potassium, on the opposite end of the reactivity series, is so reactive with water and oxygen that it is not found in pure metal form in nature. Instead it is always present as K^+ ions and when it is refined to form $K_{(s)}$ metal, it must be carefully stored in oil to prevent it from reacting violently with oxygen in the atmosphere.

For example, in a solution of dilute 0.01 M hydrochloric acid, gold, nickel, zinc and magnesium will all react to different degrees. As we learned in chapter 8, the reaction between a reactive metal and an acid will produce hydrogen gas, $H_{2(g)}$. However, how readily and vigorously the reaction occurs depends on the reactivity of the specific metal. This comparison is shown in figure 2.

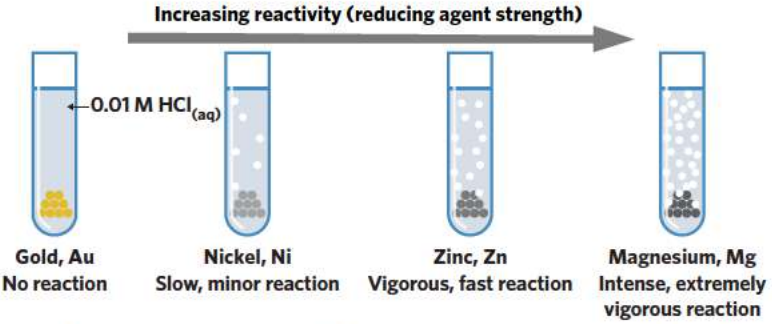


Figure 2 Metals of greater reactivity will react more vigorously with dilute acid to produce $H_{2(g)}$.

Though not shown here, the reaction between potassium and hydrochloric acid is so intense that it can be considered explosive.

In VCE Chemistry Units 3 & 4, a more comprehensive version of a reactivity series is given in the data book. This version, known as the electrochemical series, also incorporates the relative reactivity of non-metal compounds such as $Cl_{2(g)}$ and $H_2O_{(l)}$.

Highly reactive metals are strong reducing agents. Therefore, they will readily undergo oxidation.

Worked example 1

A bar of sodium (Na) metal and a bar of iron (Fe) metal are each added to a separate 500 mL beaker of deionised water. In which of the two beakers would you expect to observe a more vigorous reaction? Why?

What information is presented in the question?

$Na_{(s)}$ and $Fe_{(s)}$ are added to solutions of deionised water.

What is the question asking us to do?

Establish which metal will react more vigorously with water.

What strategy(ies) do we need in order to answer the question?

1. Determine whether $Na_{(s)}$ or $Fe_{(s)}$ is a more reactive metal.
2. Link the relative reactivities of each metal to its expected observed reaction with water.

Answer

From the reactivity series, we can see that $Na_{(s)}$ is a considerably stronger reducing agent than $Fe_{(s)}$. Therefore, $Na_{(s)}$ is a more reactive metal than $Fe_{(s)}$.

This means that the beaker containing $Na_{(s)}$ will react much more vigorously with the deionised water than the beaker containing $Fe_{(s)}$.

Metal displacement reactions 2.1.12.2

OVERVIEW

Metal displacement reactions occur when a more reactive metal displaces a less reactive metal's cations from its solution.

Metal displacement reactions occur when a more reactive metal displaces a less reactive metal from a solution of its cations.

THEORY DETAILS

So far we have only considered the reaction that would occur between a metal and oxygen, water or a solution of dilute acid. We will now investigate what happens when metals are added to solutions of other metal ions.

Consider the following redox reaction that occurs between a strip of iron metal, $Fe_{(s)}$, and a solution of copper ions, $Cu^{2+}{}_{(aq)}$.

Overall redox reaction: $Fe_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Fe^{2+}{}_{(aq)} + Cu_{(s)}$

Oxidation half-equation: $Fe_{(s)} \rightarrow Fe^{2+}{}_{(aq)} + 2e^-$

Reduction half-equation: $Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

From the reactivity series, we can see that $Fe_{(s)}$ is a more reactive metal than $Cu_{(s)}$ as it is a stronger reducing agent. This means that $Fe_{(s)}$ will readily undergo oxidation and lose its electrons, donating two electrons to the cations of the less reactive metal, $Cu^{2+}_{(aq)}$, to form $Fe^{2+}_{(aq)}$ and $Cu_{(s)}$. This kind of reaction, where a more reactive metal displaces a less reactive metal's cations from solution is known as a **metal displacement reaction**. To describe this reaction, we can say that iron has displaced copper from the solution.

Metal displacement reactions are redox reactions that occur spontaneously without needing to apply extra heat and energy and without introducing any additional chemical substances. This means this reaction can be classified as a **spontaneous redox reaction**. Whilst non-spontaneous metal displacement reactions can occur, in this course we will only consider metal displacement reactions that are spontaneous redox reactions.

The reaction between $Fe_{(s)}$ and $Cu^{2+}_{(aq)}$ is simply a reaction between the strongest reducing agent ($Fe_{(s)}$) and the strongest oxidising agent ($Cu^{2+}_{(aq)}$) present in the solution. Using the reactivity series, we can predict the outcome of metal displacement reactions by determining which chemical species is highest on the left-hand side (the strongest oxidising agent) and which chemical species is lowest on the right-hand side (the strongest reducing agent).

If we imagine a line connecting these two species, we know that for a spontaneous redox reaction to occur, this line must go downwards from left to right. As shown in figure 3, this can be likened to an 'electron slide' connecting the cation and the metal.

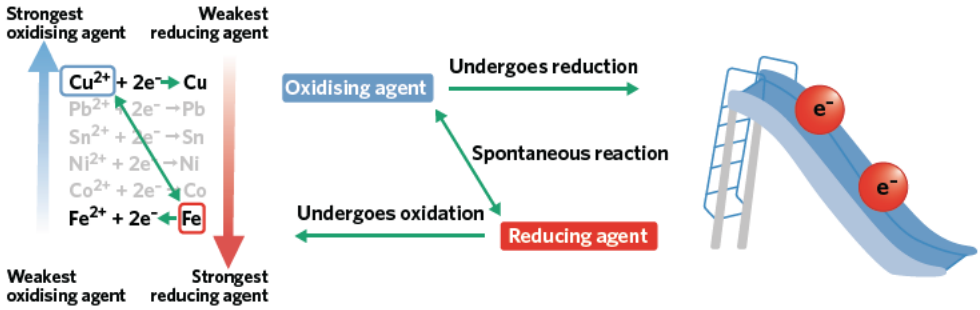


Figure 3 Spontaneous redox reactions will occur between the strongest oxidising and reducing agents present in a solution.



Another way to remember the direction of the half-reactions in a metal displacement reaction is that the lines connecting the cation and metal as well as the reduction and oxidation half-equations will form an 'S' shape if a spontaneous redox reaction occurs.



A metal ion higher on the left-hand side of the reactivity series (the oxidising agent) will react spontaneously and be displaced by a metal lower on the right-hand side of the reactivity series (the reducing agent).

Worked example 2

Write the balanced overall redox equation, if any, that would occur in each of the following scenarios. Do not include spectator ions in the overall redox equations.

a. A strip of zinc metal, $Zn_{(s)}$ is added to a solution of silver nitrate, $AgNO_{3(aq)}$.

b. A strip of zinc metal, $Zn_{(s)}$ is added to a solution of aluminium sulfate, $Al_2(SO_4)_{3(aq)}$.

---

a. A strip of zinc metal, $Zn_{(s)}$ is added to a solution of silver nitrate, $AgNO_{3(aq)}$.

What information is presented in the question?

$Zn_{(s)}$ is added to a solution of $AgNO_{3(aq)}$.

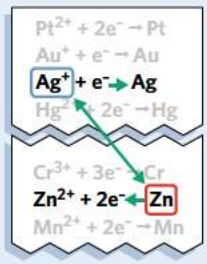
What is the question asking us to do?

Determine the balanced equation for the reaction between $Zn_{(s)}$ and $AgNO_{3(aq)}$.

What strategy(ies) do we need in order to answer the question?

1. Identify the strongest oxidising agent and reducing agent present in the beaker on the reactivity series.
2. Determine whether the reaction will be a spontaneous redox reaction.
3. Remember that the oxidising agent will undergo reduction and the reducing agent will undergo oxidation.
4. Write the individual reduction and oxidation half-equations.
5. Using the half-equations, determine the balanced overall redox reaction.

Answer



✔ **Spontaneous reaction**

We can see that the $Ag^+_{(aq)}$ ions in the solution are the strongest oxidising agent and $Zn_{(s)}$ is the strongest reducing agent. Since the line connecting the two species goes down from left to right, we know that this will be a spontaneous redox reaction. Note that the nitrate ion, $NO_3^-{}_{(aq)}$, will not take part in the reaction as it is a spectator ion.

The oxidising agent, $Ag^+_{(aq)}$ will undergo reduction according to the following half-equation.

$$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$$

The reducing agent (more reactive metal), $Zn_{(s)}$ will undergo oxidation.

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$$

In order to write the overall balanced redox equation, the reduction half-equation must be multiplied by two so that it also contains two electrons.

$$2Ag^+_{(aq)} + 2e^- \rightarrow 2Ag_{(s)}$$

Combining the two half-equations and cancelling the two electrons leaves us with the folowing balanced overall redox equation.

$$2Ag^+_{(aq)} + Zn_{(s)} \rightarrow 2Ag_{(s)} + Zn^{2+}_{(aq)}$$

---

b. A strip of zinc metal, $Zn_{(s)}$ is added to a solution of aluminium sulfate, $Al_2(SO_4)_{3(aq)}$.

What information is presented in the question?

$Zn_{(s)}$ is being added to a solution of $Al_2(SO_4)_{3(aq)}$.

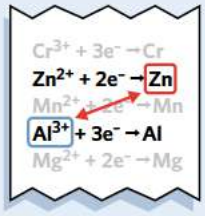
What is the question asking us to do?

Determine the balanced equation for the reaction between $Zn_{(s)}$ and $Al_2(SO_4)_{3(aq)}$.

What strategy(ies) do we need in order to answer the question?

1. Identify the strongest oxidising agent and reducing agent present in the beaker on the reactivity series.
2. Determine whether the reaction will be a spontaneous redox reaction.

Answer

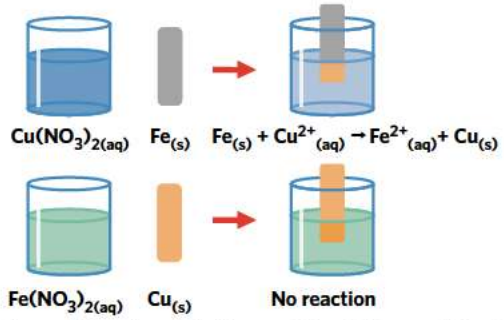


✖ **No reaction**

Since zinc is a less reactive metal than aluminium (it is a weaker reducing agent than $Al_{(s)}$), it will not be able to displace the $Al^{3+}_{(aq)}$ ions from the $Al_2(SO_4)_{3(aq)}$ solution. Therefore, there will be no observed redox reaction between $Zn_{(s)}$ and $Al_2(SO_4)_{3(aq)}$.

Note that the sulfate ion, $SO_4^{2-}{}_{(aq)}$, will not take part in any reaction as it is a spectator ion.

Experimentally, metal displacement reactions can be used to determine the relative reactivity of two metals. In the case of copper and iron, two beakers can be set up, one containing $Cu^{2+}_{(aq)}$ and $Fe_{(s)}$, the other containing $Cu_{(s)}$ and $Fe^{2+}_{(aq)}$ as shown in figure 4. Remember that the more reactive metal (the stronger reducing agent) will lose electrons and form ions in the solution. Therefore, the metal cation will accept electrons from any other more reactive solid metal atoms that are placed in the solution, resulting in these cations forming solid metal atoms themselves. However, if the solid metal is less reactive than the metal cation, there will be no reaction.



If a strip of solid iron is placed in a copper (II) nitrate solution, the iron strip will become coated with solid copper. The intense blue colour of the $Cu(NO_3)_2$ solution will fade as the $Cu^{2+}_{(aq)}$ ions are reduced to $Cu_{(s)}$ atoms.

If a strip of solid copper is placed in an iron (II) nitrate solution, there will be no reaction. There will be no solid iron deposited on the copper strip, and the pale green colour of the iron (II) nitrate solution will remain unchanged.

Figure 4 Experimental evidence of the relative reactivity of copper and iron.

The experimental set up in figure 4 reveals that iron is more reactive and a stronger reducing agent than copper. These results are confirmed by the position of these metals on the reactivity series with $Fe_{(s)}$ closer to the bottom right than $Cu_{(s)}$.

Theory summary

- The greater the reactivity of a metal, the more readily it will undergo oxidation in the presence of oxygen, water and dilute acids.
- The reactivity series of metals organises metals and their cations according to their strengths as reducing agents and oxidising agents.
  - Metal cations are found on the left-hand side of the reactivity series, increasing in oxidising strength from bottom to top.
  - Metals in their solid form are found on the right-hand side of the reactivity series, increasing in reducing strength from top to bottom.
- Metal displacement reactions occur when a more reactive metal (a stronger reducing agent) is oxidised by a less reactive metal's cations in solution.
  - The more reactive metal donates its electrons to the less reactive metal's cations, reducing them to form a solid metal deposit.
  - Having lost electrons, the more reactive metal enters the solution as cations, displacing the less reactive metal.
- To determine whether a spontaneous redox reaction will occur between a metal and a cation, the reactivity series of metals can be used.
  - First, the strongest oxidising agent and reducing agent must be identified.
  - If the line connecting the two chemical species forms a negative gradient (decreases from left to right), then the redox reaction will be spontaneous.

9C QUESTIONS

Theory review questions

Question 1

Metals that are highly reactive with oxygen, water and dilute acids are also strong _____.

A reducing agents

B oxidising agents

Question 2

$Au^{+}_{(aq)}$ is a strong _____.

A reducing agent

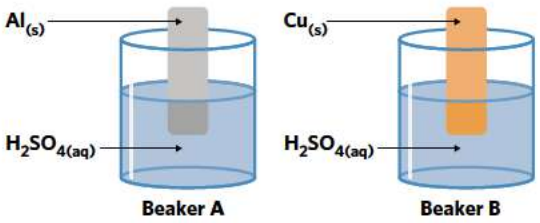
B oxidising agent

Question 3

In which of the following beakers would you expect to observe a more vigorous reaction?

A Beaker A

B Beaker B



Question 4

Reactions in which a _____ reactive metal displaces a _____ reactive metal from a solution of its cations are known as metal displacement reactions.

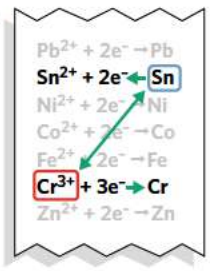
A more, less

B less, more

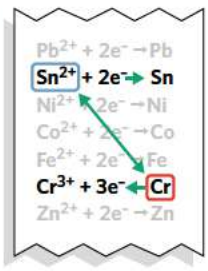
Question 5

Which of the following best represents the half-equations on the reactivity series that would occur in a metal displacement reaction between tin (Sn) and chromium (Cr)?

A



B

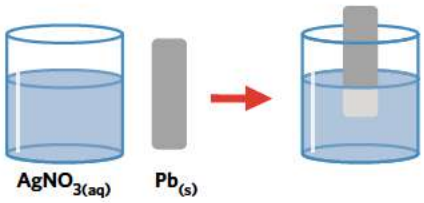


Question 6

Which of the following best describes this reaction?

A Spontaneous redox reaction

B Non-spontaneous redox reaction



Question 7

Metal displacement reactions only involve oxidation half-equations.

A True

B False

Question 8

When a strip of zinc solid, $Zn_{(s)}$, is added to a solution of $Pb^{2+}{}_{(aq)}$ ions, _____ will displace _____ from the solution.

A lead, zinc

B zinc, lead

Question 9

A highly reactive metal such as potassium, K, will react vigorously with a solution of $HNO_{3(aq)}$, but will not react with oxygen from the atmosphere.

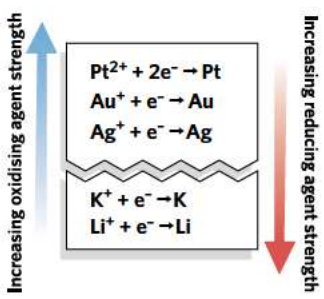
A True

B False

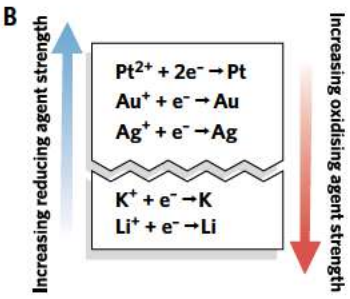
Question 10

Which of the following correctly labels the reactivity series of metals?

A



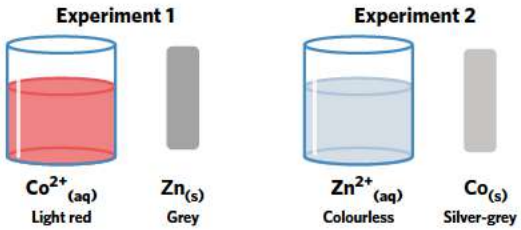
B



Deconstructed questions

Use the following information to answer questions 11–13.

A student, Aeneas, prepares solutions of $Co^{2+}_{(aq)}$ ions and $Zn^{2+}_{(aq)}$ ions. $Co^{2+}_{(aq)}$ ions form a light red solution, whereas $Zn^{2+}_{(aq)}$ ions are colourless in solution. He places a strip of grey $Zn_{(s)}$ into the $Co^{2+}_{(aq)}$ solution (experiment 1) and a strip of silver-grey $Co_{(s)}$ into the $Zn^{2+}_{(aq)}$ solution (experiment 2).



Question 11 (1 MARK)

Which of the following is a correct statement about the reactivity of zinc and cobalt?

A $Zn_{(s)}$ is a stronger oxidising agent than $Co_{(s)}$.
B $Co_{(s)}$ is a stronger oxidising agent than $Zn_{(s)}$.
C $Zn_{(s)}$ is a stronger reducing agent than $Co_{(s)}$.
D $Co_{(s)}$ is a stronger reducing agent than $Zn_{(s)}$.

Question 12 (1 MARK)

Which of the following best describes the reaction(s) that Aeneas would observe in experiments 1 and 2?

A There will be a reaction in both experiment 1 and experiment 2.
B There will be a reaction in experiment 1, but not in experiment 2.
C There will be a reaction in experiment 2, but not in experiment 1.
D There will not be a reaction in either experiment 1 or experiment 2.

Question 13 (2 MARKS)

Give the balanced oxidation and reduction half-equations for any reactions occurring in the two experiments. Hence, give the balanced overall equation(s) for these reaction(s).

Exam-style questions

Within lesson

Question 14 (1 MARK)

A strip of chromium metal is added to a solution of tin (II) cations. Which of the following gives the balanced equation for the reaction, if any, that would occur?

A $Cr^{3+}_{(aq)} + Sn^{2+}_{(aq)} \rightarrow Cr_{(s)} + Sn_{(s)}$
B $Cr_{(s)} + Sn^{2+}_{(aq)} \rightarrow Cr^{3+}_{(aq)} + Sn_{(s)}$
C $2Cr_{(s)} + 3Sn^{2+}_{(aq)} \rightarrow 2Cr^{3+}_{(aq)} + 3Sn_{(s)}$
D No spontaneous reaction would occur between $Cr_{(s)}$ and $Sn^{2+}_{(aq)}$ ions.

Question 15 (1 MARK)

Which of the following only lists chemical species that $Fe_{(s)}$ will displace from solution?

A $Zn_{(s)}$, $Mn_{(s)}$, $Ca_{(s)}$, $Mg_{(s)}$
B $Pb_{(s)}$, $Ni_{(s)}$, $Cu_{(s)}$, $Co_{(s)}$
C $Zn^{2+}_{(aq)}$, $Mn^{2+}_{(aq)}$, $Ca^{2+}_{(aq)}$, $Mg^{2+}_{(aq)}$
D $Pb^{2+}_{(aq)}$, $Ni^{2+}_{(aq)}$, $Cu^{2+}_{(aq)}$, $Co^{2+}_{(aq)}$

Question 16 (6 MARKS)

Consider the following pairs of solid metals and metal cations. If a spontaneous metal displacement reaction occurs, give its balanced overall equation. If no metal displacement reaction occurs, state 'No reaction'.

a $Pb^{2+}_{(aq)}$ and $Co_{(s)}$ (1 MARK)

b $Ag^{+}_{(aq)}$ and $Cu_{(s)}$ (1 MARK)

c $Sn^{2+}_{(aq)}$ and $Al_{(s)}$ (1 MARK)

d $Fe^{2+}_{(aq)}$ and $Cu_{(s)}$ (1 MARK)

e $Mg_{(s)}$ and $Au^{+}_{(aq)}$ (1 MARK)

f $Na^{+}_{(aq)}$ and $Zn_{(s)}$ (1 MARK)

Question 17 (1 MARK)

A shiny metal coating is formed when a strip of metal $X_{(s)}$ is added to solution of $A^{2+}_{(aq)}$ cations, but not when metal $X_{(s)}$ is added to a solution of $B^{2+}_{(aq)}$ cations. Which of the following most likely represents the identities of metals X, A and B?

	X	A	B
A	Ni	Au	Mn
B	Co	Cu	Ag
C	Al	Ca	Pb
D	Zn	Sn	Mg

Multiple lessons

Question 18 (8 MARKS)

Lausus is a chemistry student who performs the following experiments.

Beaker 1: One strip of nickel metal is added to a 750 mL solution of 1.2 M copper (II) nitrate.

Beaker 2: A second strip of nickel metal is added to a 500 mL solution of 0.90 M chromium (III) nitrate.

a Which beaker contains a higher concentration of $NO_3^-{}_{(aq)}$ ions? Include any relevant calculations in your answer. (2 MARKS)

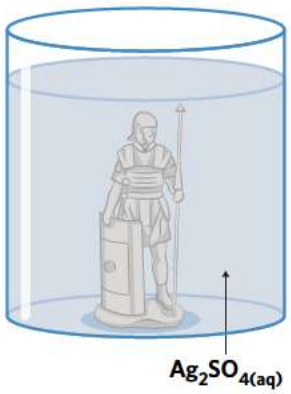
b Which beaker contains a higher number of mol of $NO_3^-{}_{(aq)}$ ions? Include any relevant calculations in your answer. (2 MARKS)

c In which beaker(s), if any, will the concentration of $Ni^{2+}_{(aq)}$ ions increase? Explain with the use of any relevant balanced overall redox equation(s). (2 MARKS)

d Will the pH of the solutions in beakers 1 or 2 change significantly as a result of these experiments? Explain. (2 MARKS)

Question 19 (9 MARKS)

When preparing a solution of $Ag_2SO_{4(aq)}$, Dido accidentally knocked her teacher's favourite figurine into the beaker – a Roman soldier made of 102 g of pure tin metal. After spending a few minutes panicking about what to do, Dido extracted the tin figurine from the $Ag_2SO_{4(aq)}$ solution, hoping that her teacher wouldn't realise anything had happened. Unfortunately for Dido, her teacher noticed immediately and as punishment asked her to answer the following questions.



a Give the balanced redox half-equations for the metal displacement reaction that occurred in the beaker. (2 MARKS)

b Hence, give the balanced overall redox equation for this reaction. (1 MARK)

c When the figurine was re-weighed, Dido found that it weighed an extra 37.4 g.

i Explain why a change in mass was observed. (2 MARKS)

ii Calculate the mass of $Ag_{(s)}$ that has been deposited on the tin figurine. (4 MARKS)

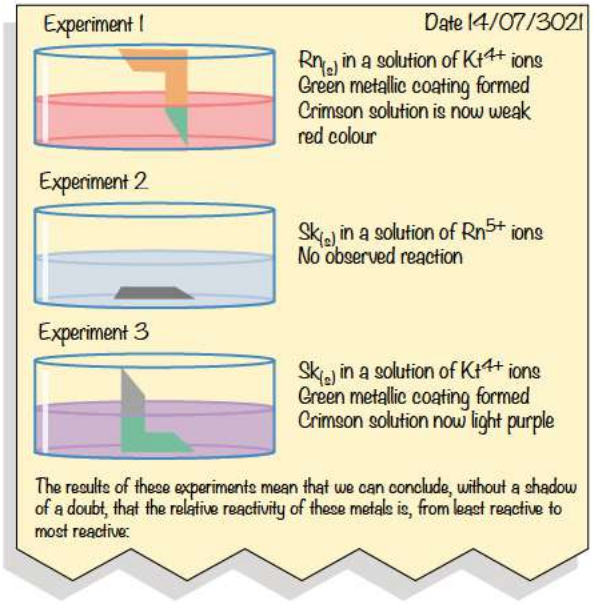
Key science skills

Question 20 (7 MARKS)

It is the year 3021 and a recent interstellar expedition to the exoplanet Sutherlandia by the chemistry syndicate Igao Industries has led to the discovery of three new metal elements.

Of course, the first thing Igao Industries does is determine the relative reactivity of these three new elements – Kostrethium (Kt), Sparkonium (Sk) and Roanorium (Rn).

An extract from the notebook of the expedition's chief scientist, Dr Komezyrov, is given.



a Give the balanced overall redox equation for the reaction that occurred in experiment 1. (2 MARKS)

b Tragically, a horde of space edroliphants ambushed the expedition's colony and Dr Komezyrov's notebook was damaged. Complete Dr Komezyrov's notes, making sure to justify your answer with reference to the results of experiments 1, 2 and 3. (5 MARKS)

Exam-style question hints

14: Metals that are strong reducing agents will undergo oxidation. **15:** More reactive metals displace less reactive metal cations from solution. **16:** More reactive metals displace less reactive metal cations from solution. **17:** Metal displacement reactions only occur if the two species form a negative gradient on the reactivity series. **18a:** $c = \frac{n}{v}$ **18b:** $n = c \times V$ **18c:** More reactive metals displace less reactive metal cations from solution. **18d:** pH is proportional to $[H^+]$. **19a:** Metal cations are oxidising agents and will undergo reduction. **19b:** Overall redox equations do not include electrons or spectator ions. **19ci:** A metal cation has been displaced from the solution. **19cii:** Solid metal is both formed and lost during this process. **20a:** More reactive metals displace less reactive metal cations from solution. **20b:** Metal displacement reactions are usually accompanied by observable changes.

9D REDOX CHEMISTRY IN THE BODY

In this lesson we will investigate the redox reactions that occur when the human body produces energy from glucose.

9A An introduction to redox reactions	9B Redox equations	9C Metal displacement reactions	9D Redox chemistry in the body

Study design dot point

- the causes and effects of a selected issue related to redox chemistry

Key knowledge unit

Redox chemistry in the body	2.1.13.1

Key terms and definitions

Cellular respiration process by which cells create energy from biochemical reactions

Glucose simple sugar molecule with molecular formula $C_6H_{12}O_6$

Adenosine triphosphate (ATP^{3-}) high-energy organic molecule used to drive cellular reactions

Nicotinamide adenine dinucleotide (NAD^+) electron carrier used in cellular respiration with reduced form NADH

Flavin adenine dinucleotide (FAD) electron carrier used in cellular respiration with reduced form $FADH_2$

Cytoplasm thick, jelly-like liquid that fills up the space within a cell

Mitochondria enclosed organelles that are the site of energy production within a cell

Metabolism process by which the body breaks down large molecules and harnesses their energy

Electron transport chain (ETC) series of protein complexes on the inner mitochondrial membrane that accept electrons from NADH and $FADH_2$ to reduce oxygen to water

Coenzyme Q mobile electron carrier in the ETC that receives electrons from NADH and $FADH_2$

Cytochrome *c* mobile electron carrier in the ETC with a central Fe^{3+} ion capable of accepting one electron

ATP synthase enzyme complex that uses protons stored in the mitochondrion's intermembrane space to drive the synthesis of ATP^{3-}

Lesson links

This lesson builds on:

- 8D Acid-base chemistry in the body

Proteins, enzymes and protons are involved in cellular respiration.

Redox chemistry in the body 2.1.13.1

OVERVIEW

Redox reactions and the movement of electrons enable body cells to harness and produce energy from food.

THEORY DETAILS

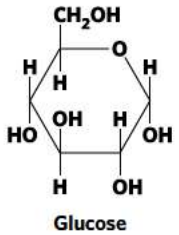
All of the cells in the human body require a constant supply of energy in order to perform their vital, life-sustaining functions. But where does this energy come from? One of the primary methods for producing energy in the human body is through the process of **cellular respiration**, which involves the oxidation of **glucose**, $C_6H_{12}O_6$. Glucose is a sugar molecule formed from the breakdown of carbohydrates such as bread, honey or potatoes and its structure is shown in figure 1. Glucose undergoes oxidation in the body to produce water, carbon dioxide and energy according to the following reaction.

$$C_6H_{12}O_{6(aq)} + 6O_{2(g)} \rightarrow 6H_2O_{(l)} + 6CO_{2(g)} + \text{energy } (ATP^{3-})$$

The production of energy from glucose, however, is not as simple as breaking the bonds in glucose and oxygen and immediately forming $H_2O_{(l)}$ and $CO_{2(g)}$ within our cells. Instead, there are many tightly controlled reactions in between these reactants and products that have to occur so that our cells can safely harness the energy from a honey sandwich.



This lesson applies the principles of redox reactions to a new context and also extends your knowledge with some of the key concepts that will arise in VCE Units 3 & 4. Don't worry about all the details of the theory, just keep an eye out for the wide variety of oxidation and reduction processes that can occur in the body!



Glucose

Figure 1 Structure of glucose.

This cellular energy is produced in the form of the energy-rich molecule, **adenosine triphosphate (ATP^{3-})** as shown in figure 2. The energy released from ATP^{3-} is used by cells to drive their molecular processes such as waste excretion and replication.

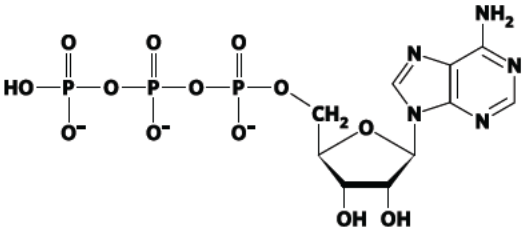


Figure 2 The structure of energy-rich adenosine triphosphate (ATP^{3-}).

As we move into redox reactions that involve organic molecules, we will expand our definitions of oxidation and reduction to simplify the redox processes that we are investigating. These definitions are given in Table 1.

Table 1 Additional definitions of redox reactions.

Oxidation	Reduction
Loss of electrons	Gain of electrons
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen

Worked example 1

The straight chain alcohol ethanol will undergo a redox reaction in the presence of an acidified potassium permanganate solution, $KMnO_{4(aq)}$, to form ethanoic acid. A schematic of this reaction is shown.



Identify whether this half-equation represents an oxidation or reduction half-reaction. Explain.

What information is presented in the question?

The structural formulas of ethanol and ethanoic acid.

What is the question asking us to do?

Identify whether the conversion of ethanol to ethanoic acid is an oxidation or reduction half-reaction and explain our answer.

What strategy(ies) do we need in order to answer the question?

1. Determine whether the reactant, ethanol, has gained or lost any hydrogen or oxygen atoms.
2. Explain whether a loss or gain of hydrogen or oxygen is indicative of an oxidation or reduction half-reaction.

Answer

Ethanol has the molecular formula C_2H_6O and ethanoic acid has the molecular formula $C_2H_4O_2$.

Therefore, ethanol has lost two hydrogen atoms and gained one oxygen atom to form ethanoic acid.

This loss of hydrogen and/or the gain of oxygen atoms indicates that ethanol has undergone an oxidation half-reaction to form ethanoic acid.

Cellular respiration also involves two crucial organic molecules that can undergo reversible oxidation and reduction. These are **nicotinamide adenine dinucleotide (NAD^+)** and **flavin adenine dinucleotide (FAD)**, which exist in either their reduced or oxidised forms as shown in figure 3. The redox conversions between NAD^+ and NADH and between FAD and $FADH_2$ are central to many of the reactions that occur in cellular respiration.

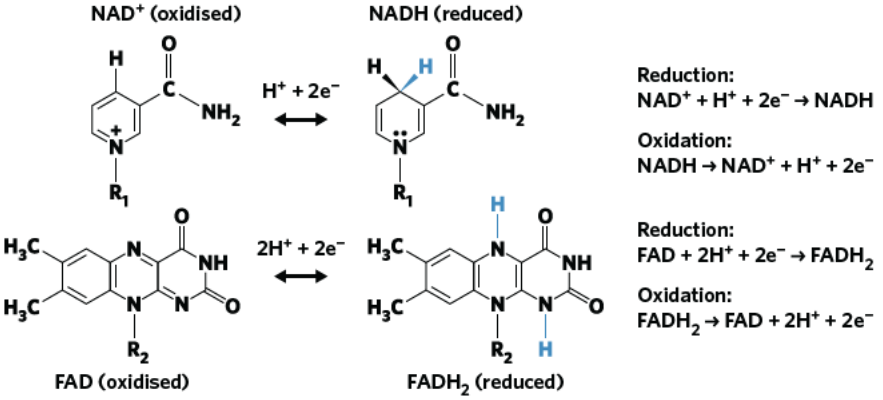


Figure 3 NADH and $FADH_2$ are the reduced forms of NAD^+ and FAD respectively. R_1 and R_2 represent large organic groups.

The metabolism of glucose

Since so many of the foods we eat contain high quantities of glucose, it is important that we can break food down relatively easily to extract the energy stored within glucose. The process by which the body breaks down a larger molecule into smaller molecules and harnesses the energy from these reactions is generally referred to as **metabolism**.

The metabolism of glucose is complex and involves over 15 biochemical reactions that take place within both the cytoplasm of the cell and within its mitochondria. The **cytoplasm** is the 'jelly-like' material that makes up much of the space within a cell, whereas the **mitochondria** are small enclosed organelles (structures) within cells that are the sites of most energy production, particularly energy production that involves the use of oxygen. For this reason, the mitochondria are often referred to as the 'powerhouses' of the cell.

The structure of a cell and one such mitochondrion (singular) is shown in figure 4. Note the inner and outer membranes, or 'leaflets' that make up the structure of the mitochondria. The space within these two membranes is referred to as the mitochondrial matrix, whereas the space between the two membranes is known as the intermembrane space (IMS).

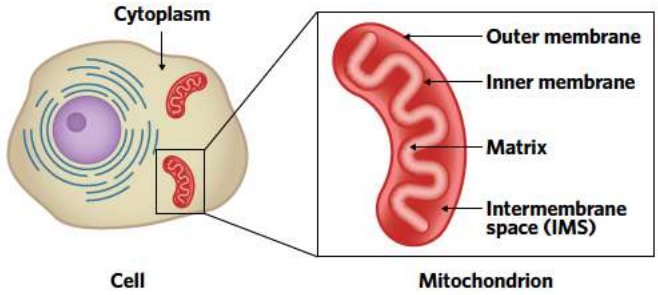


Figure 4 The cytoplasm and structure of a mitochondrion.

As shown in figure 5, the metabolism of glucose results in the production of CO_2 and some H^+ ions are released as a byproduct. It also leads to the reduction of a number of NAD^+ molecules to NADH and of FAD to $FADH_2$.

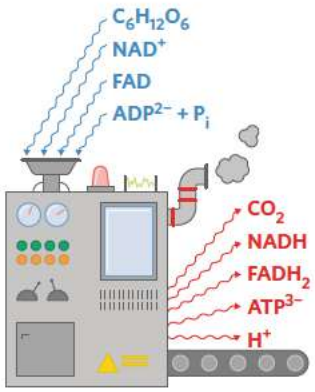


Figure 5 The inputs and outputs of glucose metabolism.

The most important products of these reactions are NADH and $FADH_2$, which are the reduced forms of NAD^+ and FAD that have gained electrons and hydrogen atoms during the oxidation of glucose.



The molecules NADH and $FADH_2$ do not appear out of nowhere or directly from the breakdown of glucose. Remember, glucose doesn't have any nitrogen! Instead, the fully oxidised forms NAD^+ and FAD are found throughout the cell and are readily available to be reduced to NADH and $FADH_2$ during cellular respiration reactions.

The electron transport chain (ETC)

After glucose has been broken down, the electrons stored in NADH and $FADH_2$ molecules are used to drive the production of lots of new ATP^{3-} molecules. This stage is referred to as the **electron transport chain (ETC)**, which uses a number of different protein complexes that are embedded in the mitochondria's inner membrane. It is in this stage that the majority of cellular energy in the form of ATP^{3-} is produced.

The NADH and $FADH_2$ molecules that have been produced are high-energy electron carriers that will undergo oxidation at the different protein complexes of the ETC. The energy from these reactions is used to drive the movement of protons and the synthesis of ATP^{3-}. A simplified view of the events of the ETC is given in figure 6.

1	At Complex I, NADH is oxidised to form NAD^+.	$NADH_{(aq)} \rightarrow NAD^+_{(aq)} + H^+_{(aq)} + 2e^-$
2	At Complex II, $FADH_2$ is oxidised to form FAD.	$FADH_{2(aq)} \rightarrow FAD_{(aq)} + 2H^+_{(aq)} + 2e^-$
3	The electrons released from NADH and $FADH_2$ are used to reduce the protein **Coenzyme Q** from its fully oxidised ubiquinone form (Q) to its fully reduced form ubiquinol (QH_2).	$Q + 2H^+_{(aq)} + 2e^- \rightarrow QH_2$
4	Coenzyme Q (as ubiquinol) transfers these electrons to Complex III and the associated protein **cytochrome c**. Losing these electrons, it reverts to its oxidised form ubiquinone.	$QH_2 \rightarrow Q + 2H^+_{(aq)} + 2e^-$
5	The central Fe^{3+} atom within cytochrome *c* accepts one electron, becoming reduced to Fe^{2+}. Two cytochrome *c* proteins transport the two electrons to Complex IV.	$2Fe^{3+}_{(aq)} + 2e^- \rightarrow 2Fe^{2+}_{(aq)}$
6	At Complex IV, these electrons from cytochrome *c* are accepted by oxygen and the Fe^{2+} ions are oxidised back to Fe^{3+} ions. Combined with two protons, oxygen is reduced to form water.	$2Fe^{2+}_{(aq)} \rightarrow 2Fe^{3+}_{(aq)} + 2e^-$ $\frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2e^- \rightarrow H_2O_{(l)}$
7	Protons are pumped into the IMS by the events of the electron transport chain and this allows ATP to be synthesised by the enzyme **ATP synthase**.	ATP^{3-} is produced

Figure 6 Summary of the steps in the electron transport chain (ETC).

Figure 7 shows the reversible reduction of ubiquinone and figure 8 depicts the protein complexes that drive the electron transport chain.

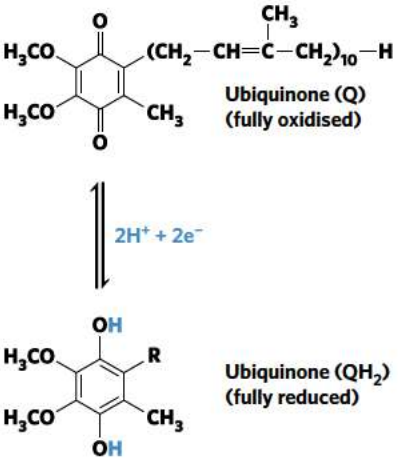


Figure 7 Ubiquinone, Q, is (reversibly) reduced to form ubiquinol, QH_2.

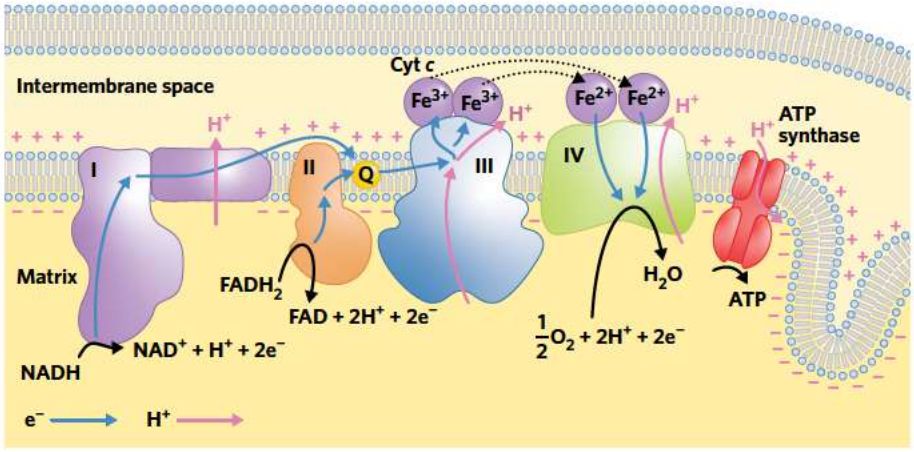


Figure 8 The protein complexes of the electron transport chain are embedded in the inner mitochondrial membrane and help drive the production of ATP^{3-}.

The electron transport chain is like a relay race. Electrons are deposited at the starting line of Complexes I and II and then transferred between Coenzyme Q, Complex III, cytochrome *c* and Complex IV. The electrons reach the finishing line when they are accepted by oxygen.

NADH and $FADH_2$ are like electron taxis that shuttle electrons around the cell. They pick up electrons in the cytoplasm and mitochondria during the metabolism of glucose and then drop them off at the electron transport chain by being reversibly oxidised and reduced.

Through the electron transport chain a large amount of ATP^{3-} can be produced relatively efficiently. This means that the mitochondria and its ETC are a very important source of cellular energy, producing as much as 90% of a cell's energy.

Mitochondrial Diseases and Inhibitors

Complex IV deficiency and Cyanide poisoning

Complex IV in the ETC is very important to the process of cellular respiration because it allows oxygen to act as the terminal electron and proton acceptor in the electron transport chain when it is reduced to form water. Complex IV deficiency is an inherited disease that results in a decreased number of functional Complex IV enzymes in the mitochondria of the patient's cells.

A more extreme mechanism for disrupting the electron transport chain is what happens when a patient is poisoned with cyanide (CN^-). Cyanide is found in different forms in many pesticides and insecticides, tobacco smoke and the smoke often produced during building fires. Cyanide poisoning from exposure to any of these sources inhibits Complex IV in the mitochondria which also prevents it from passing electrons through the electron transport chain.

A reduction in the number of Complex IV enzymes, either through a Complex IV deficiency or cyanide poisoning, disrupts the crucial final step in the electron transport chain.

These causes of mitochondrial dysfunction are summarised in figure 9.

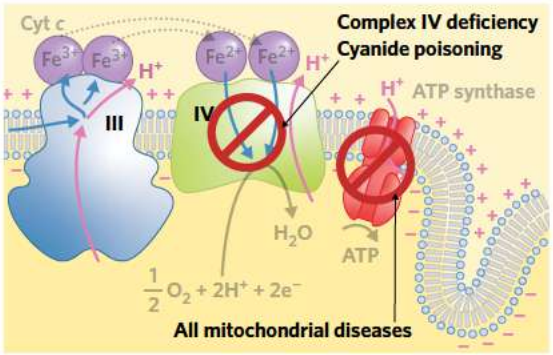


Figure 9 Mitochondrial diseases are inhibitors of the ETC.

Symptoms of mitochondrial dysfunction

What happens to the cells?

As described above, mitochondrial dysfunction will also lead to decreased ATP^{3-} production. Disrupting the ETC also tends to lead to increased amounts of acid (H^+) in the body's cells and their mitochondria. This decrease in available energy and change in H^+ concentration in cells means that they cannot continue many of their energy-intensive and pH-sensitive, but vital life-sustaining processes. For example, if a cell can no longer import nutrients and export waste products or allow the proper activity of its enzymes, it will not be able to function effectively and will instead break down and die.

What happens to the body?



Mitochondrial dysfunction results in increased acid production and decreased levels of energy in the form of ATP^{3-} produced in the body's cells. This can lead to cell death.

The inability of the body's cells to produce enough energy and the increased levels of acid circulating in the bloodstream has severe consequences for the body and its organs.

Though some mitochondrial diseases only affect specific cells or organs, most result in widespread harmful effects on the body. Mitochondrial diseases are especially damaging to the heart and the brain, as these organs require a large amount of glucose and ATP^{3-} to function properly. Since the body no longer has sufficient energy to power these organs, this means that mitochondrial diseases can lead to developmental delays in children, with the most common symptom being extreme fatigue. Over time, if the body does not produce enough energy to maintain its organ systems, mitochondrial diseases can ultimately lead to organ failure and, in severe cases, death.

The cellular and general symptoms of mitochondrial diseases and inhibitors are summarised in table 2.

Table 2 Symptoms of mitochondrial diseases and inhibitors.

Cellular Symptoms		Body Symptoms		
Low ATP^{3-}	High H^+	Dizziness	Fatigue	Death
Cell death	Enzyme dysfunction	Difficulty breathing	Developmental delays	Organ failure

Theory summary

- In the body, glucose is oxidised and broken down by the process of cellular respiration as shown in the following equation:
  - $C_6H_{12}O_{6(aq)} + 6O_{2(g)} \rightarrow 6H_2O_{(l)} + 6CO_{2(g)} + \text{energy } (ATP^{3-})$
- Oxidation can be further defined as the gain of oxygen or the loss of hydrogen.
- Reduction can be further defined as the gain of hydrogen or the loss of oxygen.
- FAD and NAD^+ are mobile electron carriers that help transfer electrons between cellular respiration reactions in the cytoplasm and mitochondria and the electron transport chain (ETC) on the inner mitochondrial membrane.
  - NAD^+ is reduced to NADH, then oxidised back to NAD^+ at Complex I of the ETC.
  - FAD is reduced to $FADH_2$, then oxidised back to FAD at Complex II of the ETC.
- The electron transport chain uses a number of different proteins and enzymes to oxidise NADH and $FADH_2$ and produce large amounts of ATP^{3-} through the enzyme ATP synthase.
- Mitochondrial diseases and inhibitors lead to decreased ATP^{3-} production and increased levels of H^+ in cells and the bloodstream.

9D QUESTIONS

Theory review questions

Question 1

Glucose is broken down by cells in a process known as _____.

A glucosal degradation

B cellular respiration

Question 2

What are the products produced during the breakdown of glucose in cells?

A H_2O, CO_2, ATP^{3-}

B O_2, ADP^{2-}, P_i

Question 3

In organic chemistry, a reduction half-reaction can be the gain of _____ atoms or the loss of _____ atoms.

A oxygen, hydrogen

B hydrogen, oxygen

Question 4

Which of the following gives the correct balanced equation for the complete reduction of flavin adenine dinucleotide (FAD)?

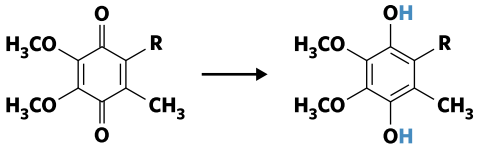
A $FAD_{(aq)} + 2H^+_{(aq)} + 2e^- \rightarrow FADH_{2(aq)}$

B $FAD_{(aq)} + H^+_{(aq)} + e^- \rightarrow FADH_{(aq)}$

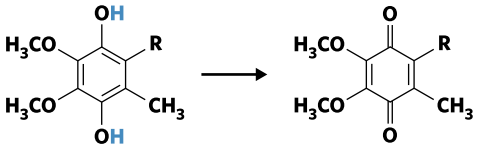
Question 5

Which of the following depicts the **reduction** of Coenzyme Q from ubiquinone to ubiquinol?

A



B



Question 6

In the cell cytoplasm and mitochondrial matrix, _____ are produced. At the electron transport chain, these molecules are oxidised to form _____.

A NADH and $FADH_2$, NAD^+ and FAD

B NAD^+ and FAD, NADH and $FADH_2$

Question 7

Which of the following represents the balanced redox half-equation for the reduction of $NAD^+_{(aq)}$?

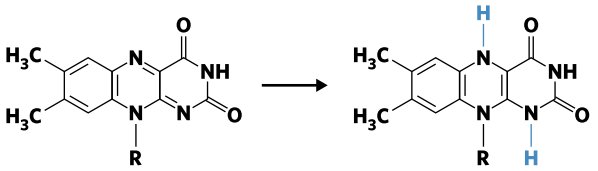
A $NAD^+_{(aq)} + 2H^+_{(aq)} + e^- \rightarrow NADH_{(aq)}$

B $NAD^+_{(aq)} + H^+_{(aq)} + 2e^- \rightarrow NADH_{(aq)}$

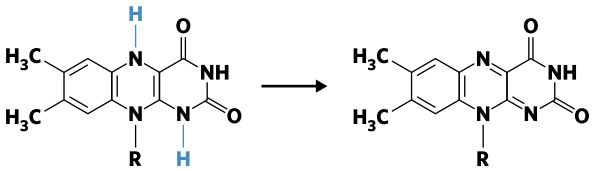
Question 8

Which of the following represents the oxidation of flavin adenine dinucleotide (FAD) as it would occur at Complex II in the electron transport chain?

A



B



Question 9

Which of the following redox reactions involving $O_{2(g)}$ occurs at Complex IV of the electron transport chain?

A $H_2O_{(l)} \rightarrow \frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2e^-$

B $\frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2e^- \rightarrow H_2O_{(l)}$

C $C_6H_{12}O_{6(aq)} + 6O_{2(g)} \rightarrow 6H_2O_{(l)} + 6CO_{2(g)}$

Question 10

Mitochondrial diseases that disrupt the electron transport chain tend to result in _____ levels of $H^+_{(aq)}$ ions and _____ levels of ATP^{3-} circulating in the bloodstream.

A increased, decreased

B decreased, increased

Deconstructed questions

Use the following information to answer questions 11–13.

Cyanide poisoning from direct exposure to insecticides, pesticides or building fires can cause severe damage to cells and tissues in the body. Cyanide interferes with the mitochondrion's electron transport chain, inhibiting the activity of Complex IV.

Question 11 (1 MARK)

Which of the given statements about cellular respiration and the electron transport chain are correct?

I The events of the electron transport chain and ATP synthase allow large amounts of ATP^{3-} to be produced.

II The electron transport chain consists of the enzymes and proteins of the inner mitochondrial membrane.

III The conversion of NADH to NAD^+ is not a redox reaction as there is no transfer of electrons involved, only the transfer of a hydrogen atom.

IV Disrupting the ETC leads to increased ATP^{3-} levels and decreased H^+ levels in cells.

V The oxidation of $FADH_2$ and NADH is coupled to the reduction of Coenzyme Q from ubiquinone (Q) to ubiquinol (QH_2).

A I, II, III

B I, III, V

C I, II, V

D III, IV

Question 12 (1 MARK)

Which of the following redox processes will cyanide poisoning have a **direct** effect on?

A The oxidation of Coenzyme Q from ubiquinol, QH_2, to ubiquinone, Q.

B The reduction of oxygen to water according to the equation $\frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2e^- \rightarrow H_2O_{(l)}$.

C The reduction of the Fe^{3+} ion within cytochrome *c* to an Fe^{2+} ion.

D The breakdown of ATP^{3-} to produce usable cellular energy.

Question 13 (3 MARKS)

Would a patient experiencing the early stages of cyanide poisoning be more likely to experience a sudden burst of energy or muscle fatigue? Explain.

Exam-style questions

Within lesson

Question 14 (1 MARK)

At Complex IV of the electron transport chain, the $Fe^{2+}_{(aq)}$ ions within two cytochrome *c* molecules are oxidised to $Fe^{3+}_{(aq)}$. Oxygen is also reduced to form water, $H_2O_{(l)}$. Which of the following gives the balanced overall equation for the redox reaction occurring at Complex IV?

A $H_2O_{(l)} + Fe^{3+}_{(aq)} \rightarrow \frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + Fe^{2+}_{(aq)}$

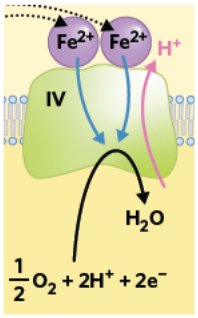
B $\frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + Fe^{2+}_{(aq)} \rightarrow H_2O_{(l)} + Fe^{3+}_{(aq)}$

C $H_2O_{(l)} + 2Fe^{3+}_{(aq)} \rightarrow \frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2Fe^{2+}_{(aq)}$

D $\frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2Fe^{2+}_{(aq)} \rightarrow H_2O_{(l)} + 2Fe^{3+}_{(aq)}$

Question 15 (5 MARKS)

Complex IV deficiency is an inherited mitochondrial disease that results in a decrease in the amount of functional Complex IV within the electron transport chain as shown.



A student knows that Complex IV is the site at which oxygen is reduced to water after electrons are delivered by the Fe^{2+} ions in cytochrome *c* proteins. She knows that the two half-reactions given take place at Complex IV.

$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^-$

$\frac{1}{2}O_{2(g)} + 2H^+_{(aq)} + 2e^- \rightarrow H_2O_{(l)}$

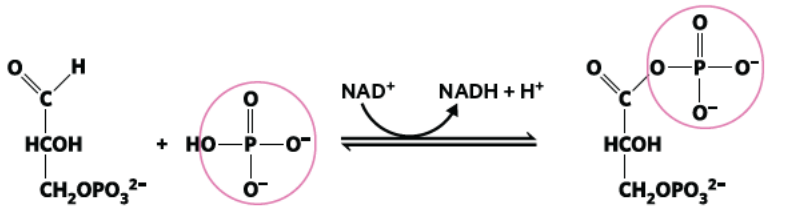
a Give the balanced overall equation for the redox reaction that occurs at Complex IV if one $O_{2(g)}$ molecule is completely reduced. (1 MARK)

b With reference to the given reactions, explain why four cytochrome *c* proteins are required to produce two molecules of water. (2 MARKS)

c Does this suggest that a patient suffering from Complex IV deficiency will be more likely to have more molecules of cytochrome *c* in the Fe^{2+} or Fe^{3+} form? Explain. (2 MARKS)

Question 16 (1 MARK)

One of the redox reactions that occurs in cellular respiration before the electron transport chain is the conversion of glyceraldehyde-3-phosphate to 1,3-bisphosphoglycerate. This reaction is coupled (occurs simultaneously) with the conversion of NAD^+ to NADH. From this information, which of the following can reasonably be concluded?



I The conversion of glyceraldehyde-3-phosphate to 1,3-bisphosphoglycerate is an oxidation process.

II The conversion of glyceraldehyde-3-phosphate to 1,3-bisphosphoglycerate is a reduction process.

III The conversion of 1,3-bisphosphoglycerate to glyceraldehyde-3-phosphate is a reduction process.

IV The conversion of 1,3-bisphosphoglycerate to glyceraldehyde-3-phosphate is an oxidation process.

V The reaction involves the transfer of electrons.

A II and IV

B II, IV and V

C I and III

D I, III and V

Question 17 (7 MARKS)

Coenzyme Q deficiency is a mitochondrial disease similar to those studied in this lesson. It leads to a decrease in the level of the important electron carrier, Coenzyme Q, in the electron transport chain. At Complex I of the ETC, NADH is converted to NAD^+ and at Complex II, $FADH_2$ is converted to FAD. At both Complex I and Complex II, Coenzyme Q as ubiquinone, Q, is also simultaneously converted to form ubiquinol, QH_2.

Note that states are not required for the equations in this question.

a Give the balanced half-equations and overall redox equation for the reaction that occurs at Complex I. (2 MARKS)

b Give the balanced half-equations and overall redox equation for the reaction that occurs at Complex II. (2 MARKS)

c Given that the reactions that occur at Complexes I and II are important parts of the electron transport chain for the later synthesis of ATP^{3-}, how will Coenzyme Q deficiency likely affect the concentrations of NADH and ATP^{3-} present in the body's cells? Explain. (3 MARKS)

Multiple lessons

Question 18 (7 MARKS)

Although the exact stoichiometry of cellular respiration is very complex, one representation of the overall equation for the oxidation of glucose is as follows [$M(C_6H_{12}O_6) = 180.0$ g mol^{-1}]. Note that states are not required for this question.

$C_6H_{12}O_6 + 6O_2 + 32ADP^{2-} + 32PO_4^{3-} + 64H^+ \rightarrow 6CO_2 + 32ATP^{3-} + 38H_2O$

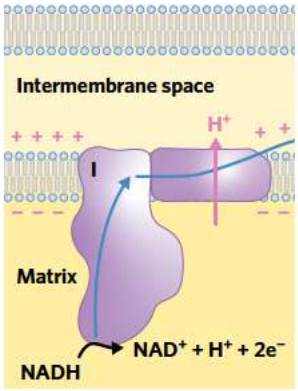
a Identify the oxidation number of the carbon atoms in glucose and carbon dioxide. Does this suggest glucose has undergone oxidation or reduction? (2 MARKS)

b Identify the oxidation number of the oxygen atoms in O_2 and in H_2O. Does this suggest oxygen has undergone oxidation or reduction? (2 MARKS)

c Consider a particular brand of jelly bean known as GlucaBeans that has recently entered the market. GlucaBeans claim to be made up of 90% pure glucose. One GlucaBean weighs 1.12 g. However, a rival jelly bean company is sceptical of their claim and performs a number of tests. One test discovered that eating five GlucaBeans resulted in the production of 0.58 mol of ATP^{3-}. Justifying your answer with a calculation, determine whether GlucaBean's claim is true. (3 MARKS)

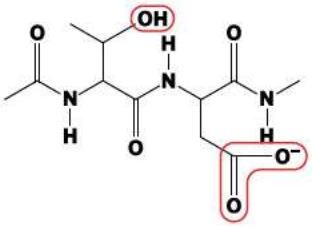
Question 19 (4 MARKS)

As a result of the reactions in the electron transport chain, a number of protons are pumped into the mitochondrion's intermembrane space. For example, at Complex I the conversion of NADH to NAD^+ drives the movement of four protons from the matrix to the IMS.



a Is this conversion of NADH to NAD^+ a half-reaction or an overall redox reaction? (1 MARK)

b Complex I consists of multiple proteins with highly complex organic structures. A small part of one such protein's structure is given.



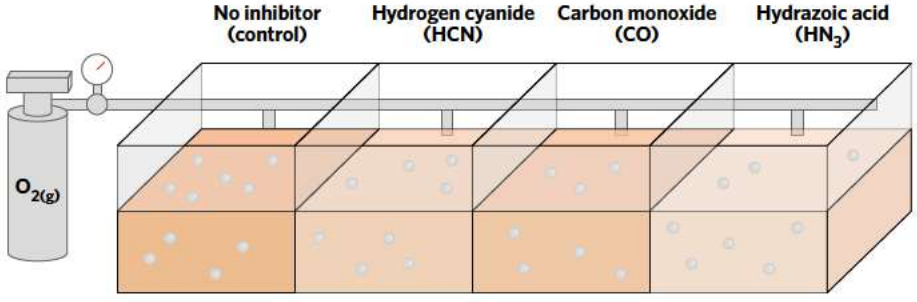
Give the names of the functional groups circled. (2 MARKS)

c Studies of the mitochondria have determined that there is a pH difference of about −1 between the mitochondrial matrix (pH ~ 8.3) and the IMS (pH ~ 7.3). What does this suggest about the relative concentrations of $H^+_{(aq)}$ in these two parts of the mitochondria? (1 MARK)

Key science skills

Question 20 (4 MARKS)

Camilla is a chemistry student who has recently learned about the redox reactions in the electron transport chain. She is interested in investigating how a certain species of microbe, *Edrolinus Chemistridium*, responds to different mitochondrial electron transport chain Complex IV inhibitors. She places four separate but identical microbial cell cultures in solutions with a different mitochondrial inhibitor, as well as a no inhibitor control. Each solution also contains 3.0 mM glucose and a pH 7.3, as well as certain ions necessary for proper cell growth and function.



Each solution contains:

- pH ~ 7.3
- $O_{2(g)}$ bubbled through
- Ions and growth factors
- 3.0 mM glucose
- *E. Chemistridium* cultures
- 1.0 mM mitochondrial inhibitor

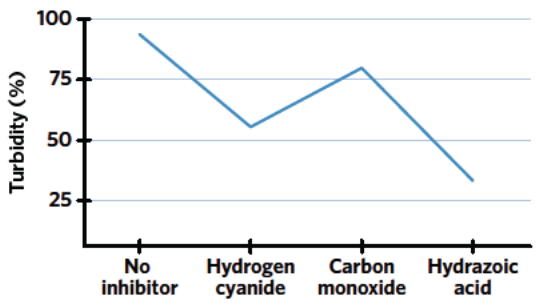
Although *Edrolinus Chemistridium* is too small to be seen without a microscope, after speaking to her biology student friends, Camilla learns that the total number of microbes in the solution is proportional to its turbidity, or 'cloudiness'. The results from her experiment are shown.

Inhibitor (1.0 mM)	Turbidity (%)
No inhibitor (control)	94
Hydrogen cyanide (HCN)	57
Carbon monoxide (CO)	78
Hydrazoic acid (HN_3)	32

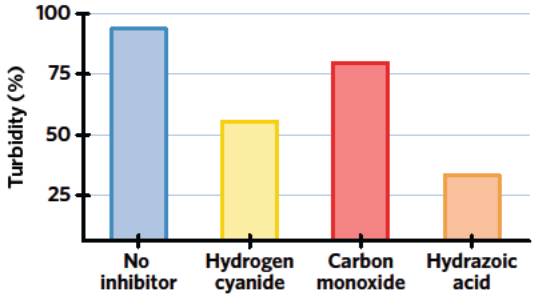
a Has Camilla collected qualitative or quantitative data in this experiment? (1 MARK)

b Which of the following would be the best method of representing the data Camilla has collected in her experiment? (1 MARK)

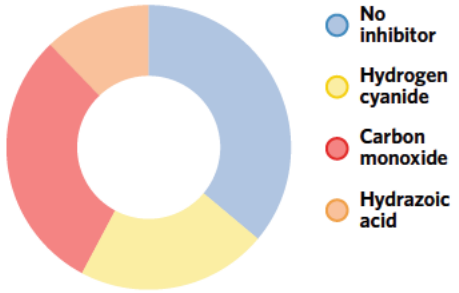
A



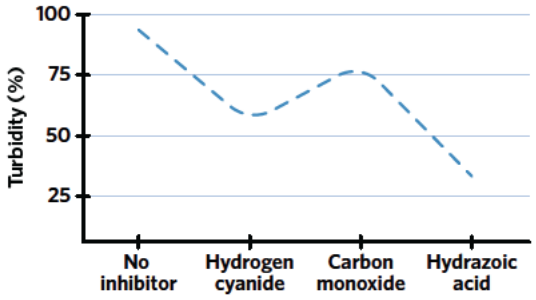
B



C



D



c If Camilla wanted to use one of the inhibitors in the experiment to stop the growth of *Edrolinus Chemistridium*, which one should she choose to best accomplish her goal? Justify your answer with reference to the results obtained. (2 MARKS)

Exam-style question hints

14: An equal number of electrons must be produced and consumed in a redox reaction. **15a:** Redox equations must be balanced for both the number of electrons and atoms of each element. **15b:** An equal number of electrons must be produced and consumed in a redox reaction. **15c** If a compound is not used, its concentration will increase. **16:** Reduction can be defined as the gain of hydrogen. **17a:** Redox equations must be balanced for both charge and number of each element. **17b:** Redox equations must be balanced for both charge and number of each element. **17c:** If a compound is not used, its concentration will increase. **18a:** Oxidation numbers in an uncharged compound sum to 0. **18b:** Free elements have an oxidation number of 0. **18c:** The amount of ATP^{3-} produced depends on the equation's stoichiometric ratios. **19a** Overall redox reactions do not show the movement of electrons. **19b:** -OH groups contain hydrogen and oxygen and $-COO^-$ groups contain carbon and oxygen. **19c:** $pH = -\log_{10}[H^+]$. **20a** Quantitative data require numerical values. **20b:** Each inhibitor has its own individual turbidity value. **20c:** A powerful mitochondrial inhibitor will prevent microbial growth.

EXPERIMENT

THE CURIOUS CASE OF THE MIXED-UP METALS

In this chapter we have investigated how the relative reactivity of metals determines how a solid metal or its cation will react in a metal displacement reaction. This can be especially useful for determining the identities of unknown metals. In this experiment, there are three unlabelled silvery-grey metals, but it is up to you (and your knowledge of redox chemistry) to determine which metal is which!

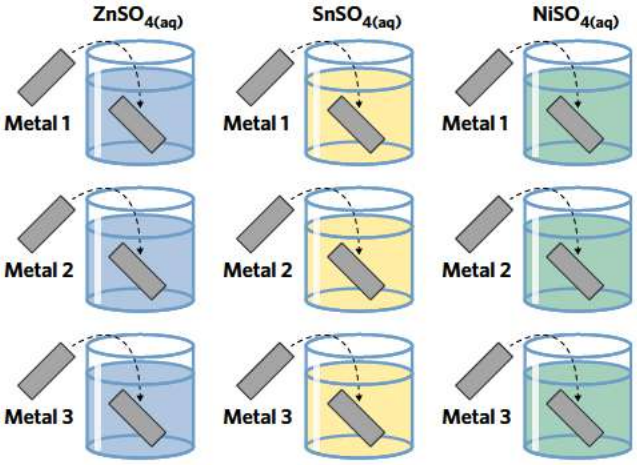
Materials

(this will be enough for two–three groups to use)

- 9 × 100 mL beakers
- 3 × Strips of zinc metal (2 cm × 5 cm)
- 3 × Strips of tin metal (2 cm × 5 cm)
- 3 × Strips of nickel metal (2 cm × 5 cm)
- 150 mL of 1.0 M $ZnSO_{4(aq)}$
- 150 mL of 1.0 M $SnSO_{4(aq)}$
- 150 mL of 1.0 M $NiSO_{4(aq)}$

Method

1. Pour 50 mL of 1.0 M $ZnSO_{4(aq)}$ solution into three different 100 mL beakers. Make sure that these beakers are labelled with $ZnSO_{4(aq)}$.
2. Repeat step 1 for the 1.0 M $SnSO_{4(aq)}$ and 1.0 M $NiSO_{4(aq)}$ solutions.
3. Label the metal strips with 1, 2 and 3.
4. Place a strip of metal 1 into each of the zinc, tin (II) and nickel (II) sulfate solutions.
5. Record any observations about the metal strips or the solution colours.
6. Repeat steps 3–4 with metal 2 and metal 3.



Results

Solution / Metal	$ZnSO_{4(aq)}$	$SnSO_{4(aq)}$	$NiSO_{4(aq)}$
Metal 1			
Metal 2			
Metal 3			

QUESTIONS

Question 1 (1 MARK)

Identify three controlled variables in this experiment.

Question 2 (1 MARK)

For this experiment, why is it important that there are three separate beakers for each of the $ZnSO_{4(aq)}$, $SnSO_{4(aq)}$ and $NiSO_{4(aq)}$ solutions instead of just one?

Question 3 (1 MARK)

One student has formulated the following table with their results from this experiment. She places a ✓ when a reaction is observed and a ✗ when no reaction is observed.

Has the student collected quantitative or qualitative data?

Solution / Metal	$ZnSO_{4(aq)}$	$SnSO_{4(aq)}$	$NiSO_{4(aq)}$
Metal 1	✗	✗	✗
Metal 2	✗	✓	✗
Metal 3	✗	✓	✓

Question 4 (3 MARKS)

Using the student's results, determine the identities of metals 1, 2 and 3. Justify your answer with reference to any relevant theory.

Question 5 (2 MARKS)

Would it be correct to state that there is a control trial in this experiment? Explain. If so, identify the control trial. If not, suggest a potential alteration to the experimental procedure that would include a control trial.

ANSWERS

1 Potential responses include:

- concentration of sulfate solutions
- volume of sulfate solutions
- sulfate anions in each ionic compound
- beaker size
- size of each metal strip

2 Separate beakers are important to prevent contamination between solutions of different ions and any ions introduced into the solution during each metal displacement reaction.

3 Qualitative data

4 [A metal displacement reaction will occur when the solid form of a more reactive metal (a stronger reducing agent) is placed in a solution of a less reactive metal's cations (a strong oxidising agent).[1]][As metal 3 displaced both the Sn^{2+} and Ni^{2+} ions from solution, it is the most reactive metal and so must be zinc.[2]][Metal 1 did not displace any ions from solution, making it the least reactive metal, tin.[3]] [This means that metal 2 must be nickel, as it displaced the Sn^{2+} ions but not the Zn^{2+} ions.[4]]

- ✓ ✗ I have described the theory of metal displacement reactions.[1]
- ✓ ✗ I have identified which metal is zinc with reference to the experimental results.[2]
- ✓ ✗ I have identified which metal is tin with reference to the experimental results.[3]
- ✓ ✗ I have identified which metal is nickel with reference to the experimental results.[4]

5 [The purpose of a control trial in an experiment is to determine whether the observed results are influenced by the independent variable using a trial for which the result is known. There is no control trial in this experiment.[1]][A potential control for this experiment would be if a fourth strip of each metal was added to a solution of highly unreactive ions (e.g. $CaSO_{4(aq)}$) or if strips of a highly unreactive metal (e.g. $Au_{(s)}$ or $Ag_{(s)}$) were added to each sulfate solution.[2]]

- ✓ ✗ I have explained whether the experiment contains a control trial.[1]
- ✓ ✗ I have suggested a possible change to the experiment to include a control trial.[2]

CHAPTER 9 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Question 1 (1 MARK)

Consider the following redox reaction between manganese (II) ions, $Mn^{2+}_{(aq)}$ and sodium bismuthate, $NaBiO_{3(s)}$ in acidic solution.

$14H^{+}_{(aq)} + 2Mn^{2+}_{(aq)} + 5NaBiO_{3(s)} \rightarrow 7H_2O_{(l)} + 2MnO_4^{-}{}_{(aq)} + 5Bi^{3+}_{(aq)} + 5Na^{+}_{(aq)}$

Which of the following correctly identifies the correct oxidising agent and reducing agent for this reaction?

	Oxidising agent	Reducing agent
A	$Mn^{2+}_{(aq)}$	$NaBiO_{3(s)}$
B	$NaBiO_{3(s)}$	$Mn^{2+}_{(aq)}$
C	$Bi^{3+}_{(aq)}$	$MnO_4^{-}{}_{(aq)}$
D	$MnO_4^{-}{}_{(aq)}$	$Bi^{3+}_{(aq)}$

Question 2 (1 MARK)

Which of the following unbalanced redox half-equations does NOT represent a reduction half-reaction?

A $O_{2(g)} \rightarrow O^{2-}_{(s)}$
B $HClO_{4(g)} \rightarrow HCl_{(g)} + H_2O_{(l)}$
C $NH_{3(aq)} \rightarrow NO_{2(g)}$
D $MnO_{2(s)} \rightarrow Mn_2O_{3(s)}$

Question 3 (1 MARK)

$Mn_2O_{3(aq)} + H_2O_{(l)} \rightarrow 2MnO_{2(s)} + 2H^{+}_{(aq)}$

Which of the following best describes the given unbalanced redox half-reaction?

A This is a reduction reaction as the oxidation number of oxygen decreases from −2 to −4.
B This is a reduction reaction as the oxidation number of manganese increases from +3 to +4.
C This is an oxidation reaction as the oxidation number of oxygen decreases from −2 to −4.
D This is an oxidation reaction as the oxidation number of manganese increases from +3 to +4.

Question 4 (1 MARK)

$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$

For this reaction to occur with the given stoichiometric ratios, how many electrons must be transferred between the aluminium atoms and the oxygen atoms?

A $3e^{-}$
B $4e^{-}$
C $6e^{-}$
D $12e^{-}$

Question 5 (1 MARK)

Consider the following reaction between iodine, $I_{2(s)}$, and lead (IV) oxide, $PbO_{2(s)}$:

$I_{2(s)} + 5PbO_{2(s)} + 8H^+_{(aq)} \rightarrow 2IO_3^-{}_{(aq)} + 5Pb^{2+}_{(aq)} + 4H_2O_{(l)}$

I $IO_3^-{}_{(aq)}$ is the conjugate oxidising agent of $I_{2(s)}$.

II This reaction is unbalanced with respect to charge.

III Electrons are lost by the $H^+_{(aq)}$ ions and accepted by $H_2O_{(l)}$.

IV $PbO_{2(s)} + 4H^+_{(aq)} + 2e^- \rightarrow Pb^{2+}_{(aq)} + 2H_2O_{(l)}$ is the balanced reduction half-equation for this reaction.

V The oxidation number of iodine increases from 0 to +5 in this reaction.

Which of the following statements about the given reaction are true?

A I, IV, V

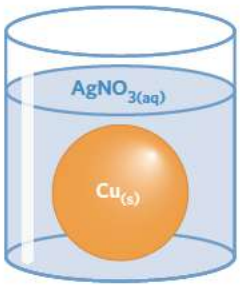
B I, II, V

C II, III, IV

D III, IV, V

Question 6 (1 MARK)

A chemistry student was interested in testing the properties of a mesmerising 99.9% solid copper sphere she had recently purchased. In order to do so, the student immersed the sphere in a solution of silver nitrate, $AgNO_{3(aq)}$.



Would a spontaneous redox reaction be expected to occur in this scenario?

A Yes, because $Cu_{(s)}$ is a stronger oxidising agent than $Ag_{(s)}$.

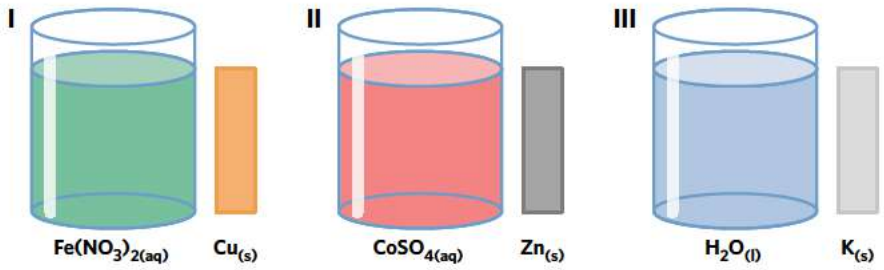
B Yes, because $Cu_{(s)}$ is a stronger reducing agent than $Ag_{(s)}$.

C No, because $Cu_{(s)}$ is a weaker oxidising agent than $Ag_{(s)}$.

D No, because $Cu_{(s)}$ is a weaker reducing agent than $Ag_{(s)}$.

Question 7 (1 MARK)

Consider the following combinations of beakers and solid metal strips.



In which beakers, if any, would a reaction be expected to occur if the metal strip was added to the beaker?

A I and II only

B II and III only

C II only

D No reaction would occur in any of the beakers.

Question 8 (1 MARK)

Of the many biochemical reactions that occur within human cells, which of the following would be considered a reduction process?

A The conversion of ubiquinone, Q, to ubiquinol, QH_2, within the electron transport chain.

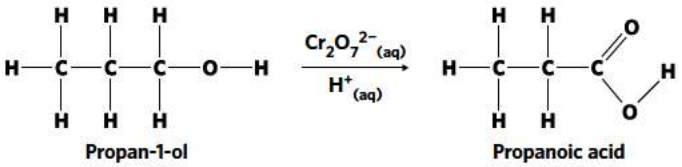
B The conversion of NADH to NAD^+ at complex I of the electron transport chain.

C The production of ATP^{3-} by the enzyme ATP synthase as usable cellular energy.

D The conversion of glucose, $C_6H_{12}O_6$, to carbon dioxide, CO_2, during cellular respiration.

Question 9 (1 MARK)

The dichromate ion, $Cr_2O_7^{2-}{}_{(aq)}$, is a powerful oxidising agent when dissolved in acidic solution. As shown, $Cr_2O_7^{2-}{}_{(aq)}$ can be used to oxidise the alcohol propan-1-ol to the carboxylic acid propanoic acid. During this redox reaction, $Cr_2O_7^{2-}{}_{(aq)}$ ions are reduced to $Cr^{3+}{}_{(aq)}$ ions.



Propan-1-ol Propanoic acid

Which of the following gives the balanced overall equation for this reaction?

A $C_3H_8O_{(aq)} + H_2O_{(l)} \rightarrow C_3H_6O_{2(aq)} + 4H^+{}_{(aq)} + 4e^-$

B $C_3H_8O_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} + 10H^+{}_{(aq)} \rightarrow C_3H_6O_{2(aq)} + 2Cr^{3+}{}_{(aq)} + 6H_2O_{(l)}$

C $2C_3H_8O_{(aq)} + 3Cr_2O_7^{2-}{}_{(aq)} + 34H^+{}_{(aq)} \rightarrow 2C_3H_6O_{2(aq)} + 6Cr^{3+}{}_{(aq)} + 12H_2O_{(l)}$

D $3C_3H_8O_{(aq)} + 2Cr_2O_7^{2-}{}_{(aq)} + 16H^+{}_{(aq)} \rightarrow 3C_3H_6O_{2(aq)} + 4Cr^{3+}{}_{(aq)} + 11H_2O_{(l)}$

Question 10 (1 MARK)

A solid iron screw is added to a 450 mL solution of silver nitrate, $AgNO_{3(aq)}$ and left for a few minutes. When the screw is removed from the solution, it is found that the beaker now also has a 0.0500 M concentration of $Fe^{2+}{}_{(aq)}$ ions.

Using this information, calculate the mass, in grams, of silver metal that must have been deposited onto the iron screw.

A 1.21 g

B 2.43 g

C 4.86 g

D 9.71 g

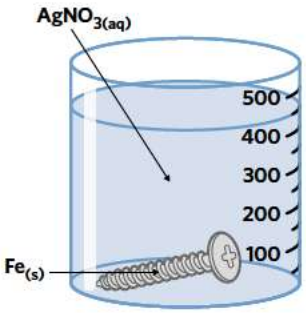


Image: Hein Nouwens/Shutterstock.com

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (7 MARKS)

Oxalic acid, $H_2C_2O_{4(aq)}$, and permanganate, $MnO_4^-{}_{(aq)}$, will react together in a redox reaction to produce $CO_{2(g)}$ and $Mn^{2+}{}_{(aq)}$ ions in an acidic solution.

a Determine the oxidation numbers for all carbon and manganese atoms in this reaction. (2 MARKS)

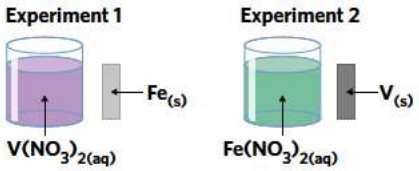
b Which chemical species is undergoing reduction and which chemical species is undergoing oxidation? Explain. (2 MARKS)

c Write balanced half-equations for the oxidation and reduction half-reactions. (2 MARKS)

d Given that one of the half-reactions involves permanganate ions and manganese ions, write the balanced overall redox equation for this reaction. (1 MARK)

Question 12 (10 MARKS)

A student is trying to determine the reactivity of the metal vanadium, $V_{(s)}$, relative to iron, $Fe_{(s)}$. To accomplish this, she sets up the two experiments shown and records her observations in the following table:



Experiment 1	No reaction observed
Experiment 2	Fizzing, solution became a darker colour, vanadium strip appears discolored/lighter

a What is the name given to the reaction occurring in Experiment 2? Is this a spontaneous or non-spontaneous redox reaction? (2 MARKS)

b Give the balanced oxidation and reduction half-equations for this reaction. Hence, give the balanced overall equation for this reaction. (2 MARKS)

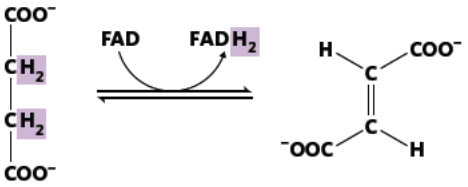
c Identify the conjugate redox pair(s) present in this reaction. (2 MARKS)

d What do the results of Experiments 1 and 2 suggest about the oxidising and/or reducing strength of vanadium metal, $V_{(s)}$, relative to iron metal, $Fe_{(s)}$? Justify your answer with reference to the experimental observations. (2 MARKS)

e If a coil of vanadium metal was added to a solution of copper (II) sulfate, $CuSO_{4(aq)}$, would a reaction be expected to occur? Explain. (2 MARKS)

Question 13 (7 MARKS)

Consider the following organic reaction that occurs during cellular respiration and the breakdown of glucose molecules. It involves the conversion of succinate, $C_4H_4O_4^{2-}{}_{(aq)}$, to fumarate, $C_4H_2O_4^{2-}{}_{(aq)}$. Coupled to this reaction is the simultaneous conversion of $FAD_{(aq)}$ to $FADH_{2(aq)}$.



Succinate Fumarate

a With reference to oxygen and/or hydrogen atoms, what evidence is there that this is a redox reaction? (1 MARK)

b Give the balanced oxidation and reduction half-equations for this reaction. Assume these reactions are occurring in a relatively acidic environment. (2 MARKS)

c Hence, give the balanced overall equation for this reaction. (1 MARK)

d Identify the chemical species that is acting as the reducing agent in this reaction. Explain. (2 MARKS)

e The $FADH_{2(aq)}$ molecules produced from this reaction are later deposited at Complex II of the mitochondrion's electron transport chain. The electrons from $FADH_{2(aq)}$ are used to convert Coenzyme Q from its Q form to its QH_2 form. Give the balanced overall equation for this redox reaction. Note that states are not required for this question. (1 MARK)

Question 14 (6 MARKS)

Consider the given unbalanced redox reaction between nitrate ions, $NO_3^-{}_{(aq)}$ and hypobromous acid, $HBrO_{(aq)}$. This reaction results in the production of dinitrogen tetroxide, $N_2O_{4(g)}$, a colourless gas and bromate ions, $BrO_3^-{}_{(aq)}$.

$NO_3^-{}_{(aq)} + HBrO_{(aq)} \rightarrow N_2O_{4(g)} + BrO_3^-{}_{(aq)}$

a Give the oxidation numbers of all atoms present in this reaction. (2 MARK)

b Determine the balanced overall equation for this redox reaction, given that it occurs in acidic solution. (2 MARKS)

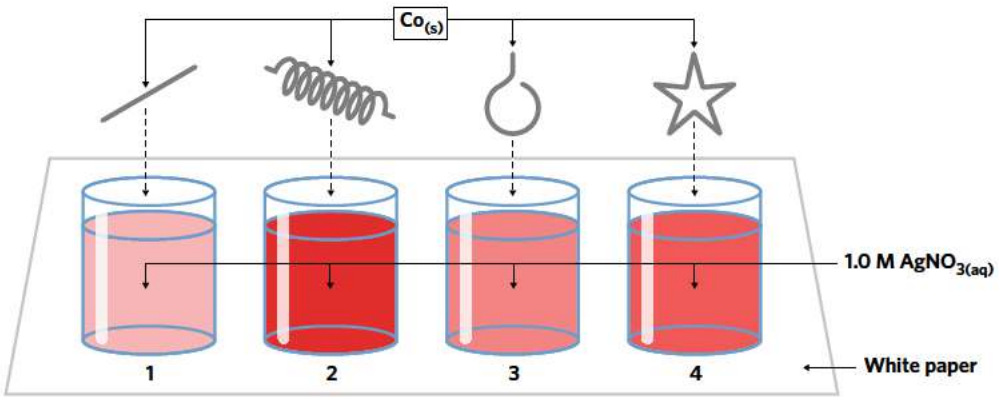
c If 2.78 mol of $N_2O_{4(g)}$ is produced in this reaction, calculate the total number of electrons that have been transferred between the nitrate ions and hypobromous acid. (2 MARKS)

KEY SCIENCE SKILLS QUESTIONS (6 MARKS)

Question 15 (6 MARKS)

Cristina is a chemistry student interested in exploring the relationship between a metal's surface area and how it behaves in a metal displacement reaction. To do this, she places a number of differently shaped pieces of cobalt wire into separate 1.0 M $AgNO_{3(aq)}$ solutions.

Cristina knows that $AgNO_{3(aq)}$ forms a colourless solution, whereas $Co^{2+}{}_{(aq)}$ ions in low concentrations form a light red solution. She measures the surface area of each piece of cobalt wire, as well as the amount of light that is absorbed by each solution after each piece of wire is left in the solution for 120 seconds. Cristina also asks her friend, Meredith, to describe and record the colour of each solution.



Beaker	Surface area (cm^2)	Absorption (%)	Description
1	7.9	24	Salmon pink
2	37.7	73	Ladybug-red
3	12.6	41	Fire-engine red
4	23.4	62	Like sunset

a Classify all the data that Cristina has collected as either qualitative or quantitative. (1 MARK)

b Cristina placed a sheet of white paper behind each beaker whenever its colour was recorded. Suggest a possible reason for this part of the experimental method. (1 MARK)

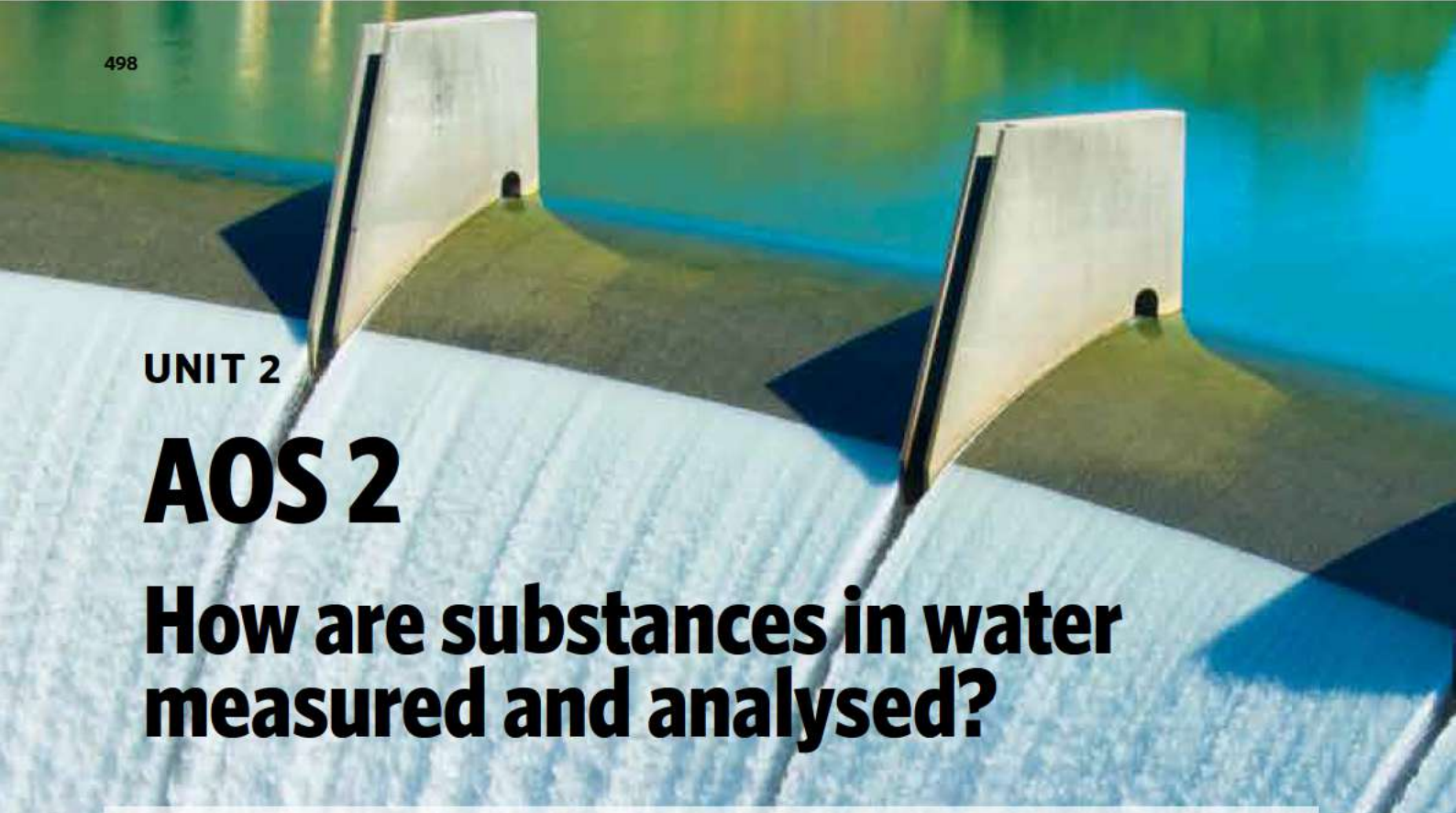
c Give the balanced overall equation for the redox reaction occurring in this experiment. (1 MARK)

d Propose a hypothesis for the relationship between the surface area of the cobalt wire and how quickly $Co_{(s)}$ is oxidised to $Co^{2+}_{(aq)}$ ions. (1 MARK)

e Meredith wants to estimate the surface area of a new piece of cobalt wire that she has taken out of her mobile phone. After replicating Cristina's experimental conditions and procedure perfectly, she writes down her results:

'The solution appears to be the shade of red similar to a ripe tomato.'

Can Meredith use this observation to estimate the surface area of the cobalt wire? Explain. (2 MARKS)



UNIT 2

AOS 2

How are substances in water measured and analysed?

In this area of study students focus on the use of analytical techniques, both in the laboratory and in the field, to measure the solubility and concentrations of solutes in water, and to analyse water samples for various solutes including chemical contaminants.

Students examine the origin and chemical nature of substances that may be present in a water supply, including contaminants, and outline sampling techniques used to assess water quality. They measure the solubility of substances in water, explore the relationship between solubility and temperature using solubility curves and learn to predict when a solute will dissolve or crystallise out of solution.

The concept of molarity is introduced and students measure concentrations of solutions using a variety of commonly used units. Students apply the principles of stoichiometry to gravimetric and volumetric analyses of aqueous solutions and water samples. Instrumental techniques include the use of colorimetry and/or UV-visible spectroscopy to estimate the concentrations of coloured species in solution, atomic absorption spectroscopy data to determine the concentration of metal ions in solution and high performance liquid chromatography data to calculate the concentration of organic compounds in solution.

Outcome 2

On completion of this unit the student should be able to measure amounts of dissolved substances in water and analyse water samples for salts, organic compounds and acids and bases.

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UNIT 2 AOS 2, CHAPTER 10

Analysis of solutions

10

10A Water on Earth

10B Water contamination

10C Analysing water using gravimetric analysis

10D Analysing water using light - part 1

10E Analysing water using light - part 2

10F Volumetric analysis

10G Chromatography

Key knowledge

- existence of water in all three states at Earth's surface including the distribution and proportion of available drinking water
- sampling protocols including equipment and sterile techniques for the analysis of water quality at various depths and locations
- options, strategies or solutions to issues related to water quality
- the definition of a chemical contaminant and an example relevant to a selected water supply
- sources of salts found in water (may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to determine the salinity of water samples
- the application of mass-mass stoichiometry to gravimetric analysis to determine the mass of a salt in a water sample
- the application of colorimetry and/or UV-visible spectroscopy, including the use of a calibration curve, to determine the concentration of coloured species (ions or complexes) in a water sample
- the application of atomic absorption spectroscopy (AAS), including the use a calibration curve, to determine the concentration of metals or metal ions in a water sample (excluding details of instrument)
- sources of organic contaminants found in water (may include dioxins, insecticides, pesticides, oil spills)
- the application of high performance liquid chromatography (HPLC) including the use of a calibration curve and retention time to determine the concentration of a soluble organic compound in a water sample (excluding details of instrument)
- sources of acids and bases found in water (may include dissolved carbon dioxide, mining activity and industrial wastes)
- volume-volume stoichiometry (solutions only) and application of volumetric analysis including the use of indicators, calculations related to preparation of standard solutions, dilution of solutions and use of acid-base titrations to determine the concentration of an acid or a base in a water sample

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10A WATER ON EARTH

In this lesson, we will be learning about the sourcing and availability of water on earth.

10A Water on Earth	10B Water contamination	10C Analysing water using gravimetric analysis	10D Analysing water using light - part 1	10E Analysing water using light - part 2	10F Volumetric analysis	10G Chromatography

Study design dot point

- existence of water in all three states at Earth's surface including the distribution and proportion of available drinking water

Key knowledge unit

The availability of water 2.2.1.1

Key terms and definitions

Potable alternative term for drinkable

Water cycle the continuous movement of water on Earth as it transitions between solid, liquid and gaseous states

Solar energy energy produced by the sun

Reservoir a water storage location

Lesson links

This lesson builds on:

- 7A Properties of water

Water can continually transition through different states.

The availability of water 2.2.1.1

OVERVIEW

The supply of clean water involves a vast number of processes and a range of infrastructure.

THEORY DETAILS

In wealthy, developed nations such as Australia, we are very lucky that the majority of the population has easy access to clean drinking water. However, there are many considerations that must be made to ensure that the water is **potable**, as well as different rules and regulations for sourcing and supplying clean drinking water. This is why it can sometimes be difficult to have clean drinking water for countries with poorer infrastructure.

Historically, aqueducts were used to transport clean water over long distances. The first aqueducts were produced in the Roman Empire in the year 312 BCE, with an example of such a structure shown in figure 1.



Image: KarSol/Shutterstock.com

Figure 1 Aqueduct of Segovia, Spain.

Since then, advances in technology have allowed us to develop better systems to supply everyone with clean water.

The water cycle

The **water cycle** describes the movements of water between states as a gas, liquid and a solid through natural processes. This water is continuously moving between the land, the ocean, the atmosphere, underground, and even passing through various life forms. This is summarised in figure 2.

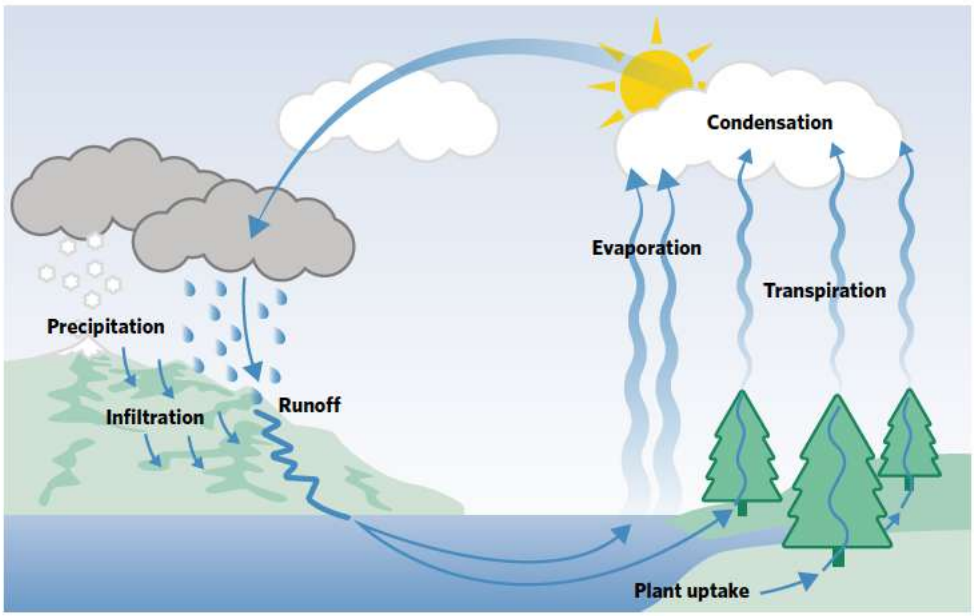


Figure 2 The water cycle.

As we can see in figure 2, the main stages of the water cycle include evaporation, transpiration, condensation, precipitation and run-off. **Solar energy** provided by the sun is used to break the intermolecular bonds between water molecules in a liquid state, allowing water to evaporate. The accumulation of gaseous water droplets results in the formation of clouds and is referred to as condensation, which subsequently allows for precipitation to eventually occur.

The water molecules on Earth and in the atmosphere are continuously moving between the different states based on the conditions they are exposed to at any given moment. To put that in perspective, it is estimated that there is approximately 1.386 billion km³ of water on Earth – that's approximately equal to filling the MCG 81 billion times – and this vast amount of water is continuously transitioning between different states.

Water can be separated into a number of categories based on the amount of salt it contains, as shown in figure 3. These include:

- Seawater – water containing high concentrations of dissolved salts found in seas and oceans.
- Brackish water – water with more salt than freshwater, but less than seawater. This is found in estuaries, where rivers meet oceans.
- Freshwater – water containing less dissolved salts than seawater and brackish water. Freshwater is found in rivers, lakes and streams. While potable water has low salt content, not all freshwater is potable water.

The human body is not effective at filtering out salt, therefore the water we consume needs to have a very low salt content. Unfortunately, the majority of the water found on earth contains salt, as shown in figure 4.

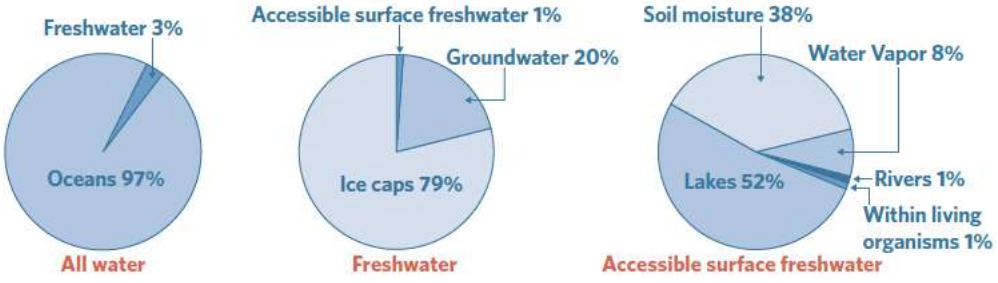


Figure 4 Distribution of water on Earth.

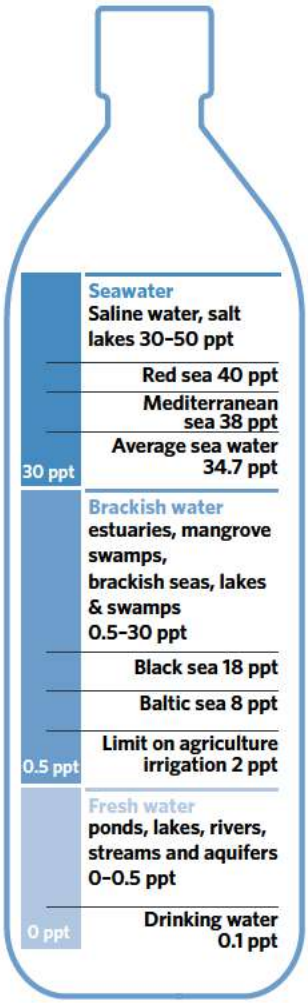


Figure 3 Types of water.

As we can see, approximately 3% of water on Earth is freshwater, and therefore more likely to be drinkable. However, most of this freshwater is trapped in ice caps, glaciers or in groundwater, making it hard to access.

The most readily accessible source of water on earth is surface freshwater, including lakes, rivers and springs. The water supplying the city of Melbourne primarily relies on **reservoirs** and catchments. The steps taken in supplying Melbourne with water include:

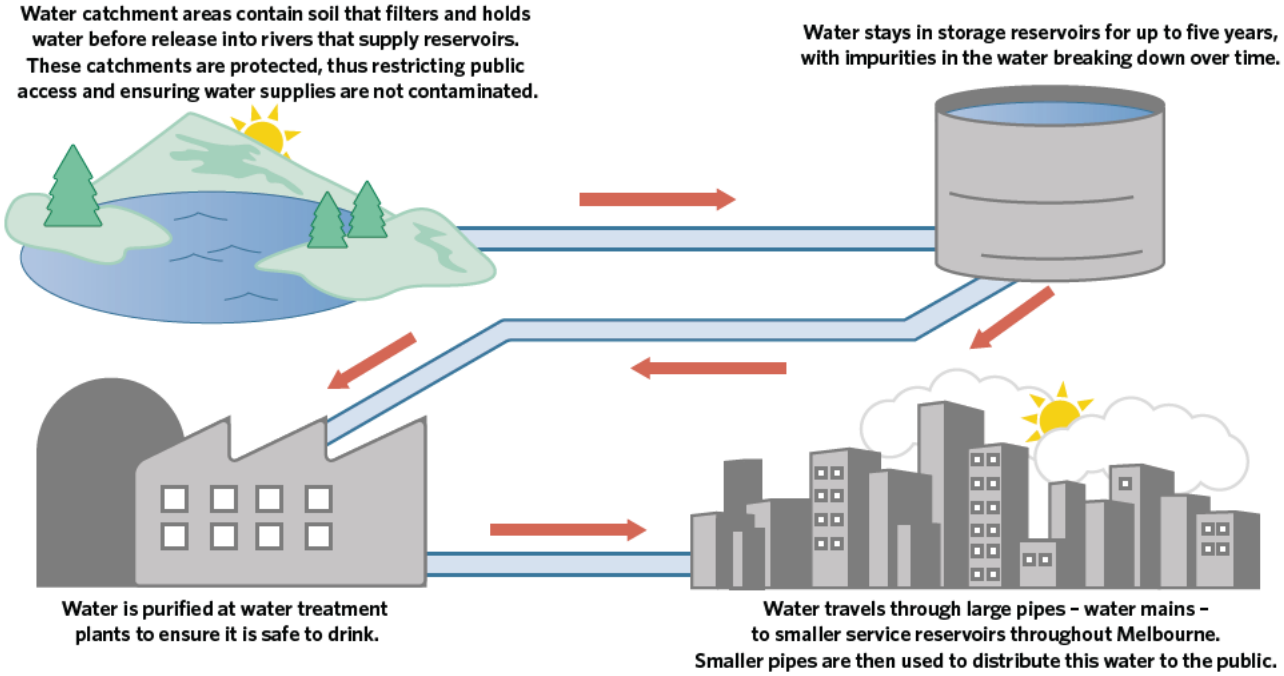


Figure 5 Sourcing and supplying potable water to the city of Melbourne.

Although figure 5 shows the primary method of sourcing potable water, there are other methods available. These include:

- Extracting groundwater using bores.
- Obtaining water directly from rivers or lakes – this is not the same as from a reservoir, as a reservoir is an often manmade lake with the primary purpose of supplying potable water.
- Processing seawater at desalination plants.

Groundwater has historically been used as a source of drinkable water by indigenous Australians. Indigenous tribes possess extensive knowledge of the usefulness of groundwater, and this forms a significant part of some indigenous cultures. It is also known that groundwater features in many cultural practices and songs of indigenous peoples, which often describe the locations, quantities and management techniques of groundwater resources.

Different sources of water possess different risks of contamination depending on the environment that they are found in. However, the thorough treatment regimen of mains water makes it the lowest risk for contamination, and therefore safest for public consumption.

Additionally, Melbourne Water also operates a number of water treatment plants to produce recycled water from wastewater. However, this water is intended for uses such as washing and garden use, and is not considered safe to drink due to potential contamination remnants.

Theory summary

- Water exists in different forms depending on the environment.
- Not all forms of water are drinkable.
- Salinity is a major determinant of whether water is safe for human consumption.
- Many different processes are required to ensure that water is safe to consume.

10A QUESTIONS

Theory review questions

Question 1

Water is

A stationary.

B constantly moving through the water cycle.

Question 2

Seawater is unable to be consumed by humans because

A it is contaminated with industrial waste.

B it contains too much salt for human consumption.

Question 3

The primary water source for Melbourne is

A natural or artificial lakes acting as reservoirs.

B water tanks located at each person's house.

Question 4

Because freshwater comes straight from natural sources, it is safe to drink.

A True

B False

Question 5

The majority of the available water on Earth is

A fresh water.

B salt water.

Deconstructed questions

Use the following information to answer questions 6–8.

Deforestation has not only caused a loss of habitat to many different animals, but has also disrupted the availability of water in many different areas on Earth.

Question 6 (1 MARK)

The process where trees and plants absorb water through their roots and release it into the atmosphere is known as

A condensation.

B purification.

C precipitation.

D transpiration.

Question 7 (1 MARK)

The process that takes place when water droplets accumulate in clouds is known as

A condensation.

B purification.

C precipitation.

D transpiration.

Question 8 (2 MARKS)

Areas in southern Western Australia have experienced periods of minimal rainfall. This began to occur just after bushland was cleared to make way for other developments. Explain why the rainfall in these areas has decreased.

Exam-style questions

Within lesson

Question 9 (7 MARKS)

Water is a very important resource that is used for a variety of different activities, from agriculture to basic human survival.

a What is the main difference in the contents of seawater, brackish water and freshwater? (1 MARK)

b Which of the types of water in question 9a are considered as safe for human consumption? Explain. (3 MARKS)

c The most common source of water for public consumption is state-run reservoirs, which are found in restricted zones.

i Suggest one reason why it may be important to prevent public access to reservoirs. (1 MARK)

ii How is water transferred from reservoirs to people's houses? (2 MARKS)

Multiple lessons

Question 10 (6 MARKS)

Seawater often contains ionic compounds such as NaCl. This is why, when swimming, the water can be described as tasting 'salty'.

a Consider a child at the beach collecting water for their sandcastle. Upon scooping up a sample of seawater into a bucket, the water seems clear and there is no distinguishable trace of NaCl. Explain this observation. (3 MARKS)

b Students at Rye beach took a sample of the seawater to school to see how 'salty' the water was. If it was shown that the concentration of NaCl was 12 g/L, how many mol of NaCl was present in a 250 mL sample of the seawater? (2 MARKS)

c If the acceptable concentration of salt in human drinking water is 0.008 M, would the sample above be considered safe to drink? (1 MARK)

Key science skills

Question 11 (4 MARKS)

Segnish wanted to test the knowledge of the water cycle of students in her class, who had recently learned about this concept in chemistry. Segnish used an unannotated water cycle diagram to identify the level of understanding of the students, who she assumed had all been in the class. In the test, the diagram was given to three different people, including two people that had attended the class about the water cycle and one person who had not.

a What is the dependent variable in this experiment? (1 MARK)

b Comment on the validity of the experiment. (2 MARKS)

c Identify one change that needs to be made to improve the validity of the experiment. (1 MARK)

Exam-style question hints

9a: Different water sources differ in the concentration of their contents. **9b:** Humans cannot efficiently remove salt from their body. **9c i:** Water supplies must be kept clean. **9c ii:** Houses are interconnected through a water network. **10a:** The properties of water allow it to interact with ionic substances differently. **10b:** The concentration of a substance doesn't change depending on the volume of the sample. **10c:** When comparing solutions, it is best to do so using the same units. **11a:** The dependent variable is the variable that is affected by the independent variable. **11b:** Validity describes whether the experiment is able to address the aim of the research. **11c:** For an experiment to show valid results, all variables other than the independent and dependent variables need to be held constant.

10B WATER CONTAMINATION

In this lesson, we will be learning about water contaminants that can affect the quality of a sample of water.

10A Water on Earth | 10B Water contamination | 10C Analysing water using gravimetric analysis | 10D Analysing water using light - part 1 | 10E Analysing water using light - part 2 | 10F Volumetric analysis | 10G Chromatography

Study design dot points

- sampling protocols including equipment and sterile techniques for the analysis of water quality at various depths and locations
- options, strategies or solutions to issues related to water quality
- the definition of a chemical contaminant and an example relevant to a selected water supply
- sources of organic contaminants found in water (may include dioxins, insecticides, pesticides, oil spills)
- sources of salts found in water (may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to determine the salinity of water samples
- sources of acids and bases found in water (may include dissolved carbon dioxide, mining activity and industrial wastes)

Key knowledge units

Water quality analysis	2.2.2.1 & 2.3.7.1
Water contamination	2.2.3.1 & 2.2.14.1
Sources of salt in water	2.2.8.1
Organic contaminants of water	2.2.12.1

Key terms and definitions

Contaminant substance that makes another substance impure and is usually toxic

Heavy metals metals with relatively high atomic masses

Eutrophication richness of nutrients in water system leading to the growth of plant-life

Hardness of water amount of calcium and magnesium dissolved in water

Bioaccumulation gradual accumulation of substances within an organism

Organometallic compounds organic compounds containing at least one metallic element in its structure, with a carbon-metal bond

Lesson links

This lesson builds on:

- 10A Water on Earth

There are different strategies used to ensure the quality of water.

Water quality analysis 2.2.2.1 & 2.3.7.1

OVERVIEW

The Environment Protection Authority enforces strict analysis protocols to ensure drinking water supplied to the population is safe.

THEORY DETAILS

As we saw in lesson 10A, there is a large number of different water sources on Earth, however, only a very limited number of these are actually safe for human consumption. Therefore, when providing the public with water, it is essential to ensure that the water does not possess dangerous levels of **contaminants**, which are substances that affect the purity of water. Failure to safely manage water contaminants could either be immediately harmful to the population, or could result in chronic illness over time. In order to ensure the water quality is upheld in Victoria, systematic water quality monitoring is carried out by the Environment Protection Authority (EPA) Victoria.

In order to assess the quality of a water supply, water sampling has to be carried out. This testing is carried out regularly, as it is possible for contamination to occur at any point in time, however, it is especially important at times when contamination may be expected. In cases of bushfires, for example, water supplies such as reservoirs can become contaminated with pollutants such as ash, debris, the remains of animals or aerial fire retardants as shown in figure 1, all of which need to be treated before human consumption can be permitted.



Image: rck_953/Shutterstock.com

Figure 1 Aerial fire retardant.

Water quality assessments involve taking water samples and putting these through a series of tests, in order to compare the results of the sample with the standard results seen in a water sample safe for human consumption. There are various methods of carrying out this testing, depending on if the researcher is conducting the test from a boat, a bridge, or on the shoreline. All of these tests use the same principles to determine water quality.

When a researcher is collecting a sample from a dam or reservoir as highlighted in figure 2, it is important that they collect samples from a number of different depths, as the composition of the water can vary depending on how deep in the water source the sample is taken. Additionally, when collecting subsequent water samples to observe how the composition of the water source has changed over time, researchers must continue to use the same depths previously used to collect samples, to ensure that any changes that are measured are accurate for that part of the water. Otherwise, due to the varying composition of the water, this may produce misleading results for the researchers.



Image: Ivan Chudakov/Shutterstock.com

Figure 2 EPA worker taking a water sample.

Another method of assessing the composition of water is through measuring the temperature at different depths in the water. If the temperature doesn't change at different depths it can be assumed that the water is at least reasonably well mixed, and so concentrations of chemicals will be approximately similar across depths.

Additionally, it's also important to collect samples from many different locations at a water source, not just at different depths, to ensure that a comprehensive set of samples has been obtained. Sometimes, as we can see very different levels of water contamination taken from different samples of the same river or lake as highlighted in figure 3.



Image: Alena Brozova/Shutterstock.com

Figure 3 Potential samples of water taken from the same source.

As we know, for the sake of good scientific method, it's absolutely essential that researchers repeat experiments not only at the same depths and locations, but also using the same methods and equipment to ensure consistency. Protocols determining sampling technique include:

- The container must not be reactive with the sample.
- The container must be rinsed before taking another sample.
- Water used to clean the container should be tested to ensure it has not been contaminated.
- Ensure the same chemicals are being analysed in each test.
- Ensure the use of the same equipment.
- Gather the required sample size.
- Ensure the same method is used for producing quantitative results of water quality.
- Ensure that labelling, storage and transport are standardised.



It is important to be familiar with water sampling protocol and why certain measures are taken to improve the accuracy and reliability of results.

Currently, Yarra Valley Water takes 7000 samples of water every year, with 22 000 tests being conducted with these samples. These samples are taken from 34 different water quality zones. The parameters that Yarra Valley Water uses to monitor the quality of the water include:

- Microbiological parameters – assessment for the presence of E.coli, a bacterial strain commonly found in the gastrointestinal tract and faecal matter of humans and other animals, which can cause severe illness.
- Chemical – assessment for trihalomethanes (THM), a by-product of water chlorination for disinfection, which requires control due to health consequences.
- Turbidity – a measure of how clear water is, indicates the amount of suspended solids in the water as shown in figure 4.

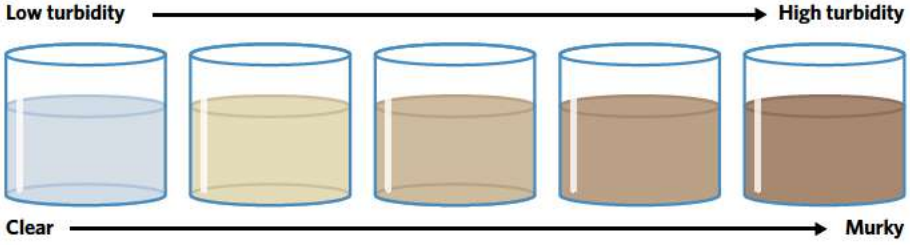


Figure 4 Water turbidity.

Some other parameters that require consideration, as determined by Melbourne Water, include:

- Concentration of Na^+, Ca^{2+}, K^+, Fe^{3+}, Cl^-.
- Nutrients such as nitrogen and phosphorus which are essential for plant and animal growth, however excessive quantities can promote growth of blue-green algae which can be toxic to the ecosystem.
- pH.
- Water hardness – corresponds to concentration of dissolved ions.
- Dissolved oxygen.
- Chlorophyll-A which is used to measure amount of algae in water.

All of these are factors that need to be closely monitored and maintained at acceptable levels deemed healthy for human consumption.

Water contamination 2.2.3.1 & 2.2.14.1

OVERVIEW

Contamination of water supplies presents a serious risk to both the environment and to the community, and comes from various sources.

THEORY DETAILS

A chemical contaminant is a chemical that is considered toxic to plants or animals that can be found in water systems. The toxicity of these substances requires the monitoring of water sources to ensure that these substances will not pose a danger to the environment or the community.

Water contamination can come from a variety of sources. As highlighted in the topic on water quality analysis, there are many programs dedicated to ensuring that the water being provided to the Australian public is safe to drink. However, it is also important to understand what exactly the danger is, why it is dangerous, and where it comes from.

Human activity is a major contributor to the introduction of chemical contaminants into water. Agricultural practices can introduce pollutants into the water while large scale industrial activities can increase the acidity of water or even cause acid rain. For example, the run-off of acidic water from metal or coal mines, also known as acid mine drainage, causes a decrease in the pH of water. The main contributor to acid mine drainage is pyrite (FeS_2) oxidation, which takes place when the mineral is exposed to air and water as shown.

$$4FeS_{2(s)} + 15O_{2(g)} + 14H_2O_{(l)} \rightarrow 4Fe(OH)_{3(s)} + 8H_2SO_{4(aq)}$$

The overall release of hydrogen ions from the reaction causes the water to become more acidic and the iron(III) hydroxide solid gives acid mine drainage waters an orange appearance, as shown in figure 5.

The increased acidity of the water makes it less suitable to support aquatic life, leading to a reduction in the number of aquatic animals and diversity of species in water supplies. There are currently efforts being made to monitor acid mine drainage and neutralise these waters to minimise the effect on aquatic wildlife.



Image: Jose Arcos Aguilar/Shutterstock.com

Figure 5 Tinto river with iron(III) hydroxide contamination.

Sources of salt in water 2.2.8.1

OVERVIEW

Salts in water can exist in a number of different forms and come from a number of different sources.

THEORY DETAILS

The presence of salts in water is a form of chemical contamination, and can come from a number of different sources.

One potential source of chemical contamination is as a result of natural processes. As discussed in lesson 10A, the water cycle involves the movement of water between different locations and states around the Earth. In this process, water runs through soil and rocks, dissolving solid mineral deposits and transporting these into freshwater sources, including lakes, rivers and reservoirs, which may not be suitable for human consumption.

When assessing for water quality, researchers will take note of the hardness of water. This is because hard water indicates that the sample contains many metallic ions, and thus requires a large amount of soap to produce a lather or a froth. Soap lather is produced from the production of stearate ions ($C_{17}H_{35}COO^-$) from sodium stearate molecules ($C_{17}H_{35}COONa$) in soap, which can then act as a dirt remover and produce a soap lather. As stearate ions are anions, these can interact with metal cations present in hard water. The interaction between stearate anions and metal cations can form a separate insoluble product also known as scum and thereby contaminants from hard water as shown in figure 6. As a consequence of this, the amount of lather produced in hard water is reduced, and soap is less effective as a dirt remover.

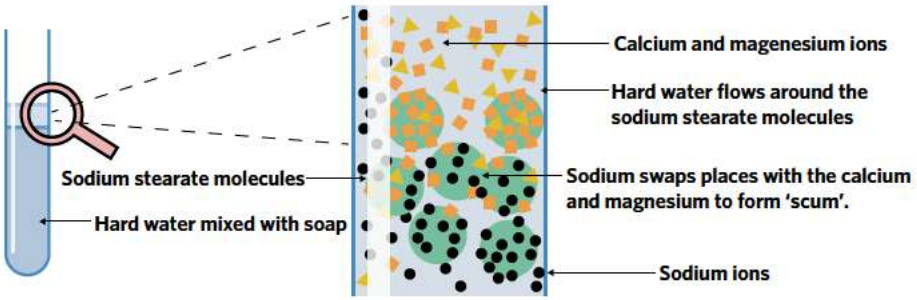


Figure 6 The action of soap in removing contaminants in hard water.

Heavy metals

It is essential to monitor the concentration of heavy metals in water systems. These metals are known as heavy metals as they have a high density, which causes them to be highly toxic. As these metals have a very high density, this causes them to interact with proteins and enzymes and consequently affecting the function of these molecules. This can have a harmful effect when it affects the structures of proteins essential in the functioning of the human body. For example high concentrations of heavy metals such as mercury and lead in the body has been seen to cause cancer, organ and nervous system damage and, in some cases, death.

Heavy metals found in the water supply have a number of different sources. These range from geological sources, with heavy metals produced naturally in the Earth's crust, to run-off from industrial processes. It is even possible to get heavy metal contamination from atmospheric sources, such as when the combustion of fuels containing heavy metals takes place, ions can be released into the atmosphere, reacting with water molecules and entering the water system. However, heavy metal contamination is most often found in areas surrounding mines, foundries and smelters, which are all locations of intense industrial activity.

The presence of heavy metals in water systems leads to a process known as **bioaccumulation**. This refers to the increasing concentration of heavy metals in the tissue of animals as one progresses up the food chain as shown in figure 7.

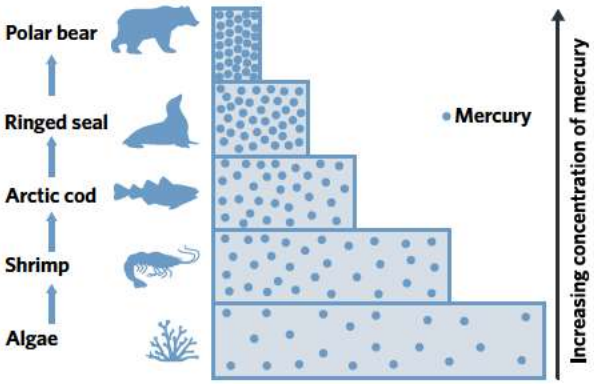


Figure 7 Bioaccumulation of heavy metals.

This takes place because, as species lower in the food chain ingest heavy metals found in water, the consumption of these lower species by higher species in the food chain will cause the heavy metals from the lower species to enter their bodies, thus increasing the concentration of heavy metals every time they eat another organism containing heavy metals. Since heavy metals are not biodegradable, and will stay in the environment they are found



Heavy metals are not biodegradable, and will persist in their environment, leading to bioaccumulation.

in, the concentration will continue to increase as consumption continues, and those organisms highest on the food chain will thus be consuming species with higher heavy metal concentrations. Consequently, those at the highest point of the food chain will be continuously consuming animals already with high heavy metal concentrations, leading to those at the top having the greatest amount of heavy metals.

Heavy metal levels in water supplies are closely monitored, and thus in order to make a water source permissible for human consumption, heavy metal contamination must be removed. According to the Australian Drinking Water Guidelines, the concentrations of heavy metals in water as outlined in table 1 are the maximum concentrations considered safe.

Table 1 Safe concentrations of heavy metals in water.

Heavy metal	Maximum concentration (mg L^{-1})
Arsenic	0.01
Cadmium	0.002
Copper	2
Chromium (as Cr (VI))	005
Lead	0.01
Mercury	0.001



Units 1 & 2 Chemistry students are not required to memorise the safe concentrations of heavy metals in water – we just want you to become familiar with the concept.

If water supplies contain heavy metal concentrations greater than the values shown in table 1, those ions must be removed from the solution. One method of doing this is through precipitating heavy metal ions out of the solution through the introduction of appropriate reagents that would produce such a reaction. This results in a solid precipitate containing the heavy metal ion, now separate from the solution and easily filtered out. Additionally, heavy metals are toxic to organisms when found as cations in a solution, however as an uncharged precipitate able to be filtered out they are no longer considered toxic.

Heavy metals can be found in two toxic forms – as a cation dissolved in a solution, and as **organometallic compounds**. Organometallic compounds are substances containing at least one carbon-metal bond. Organometallic compounds are a very large group of compounds, found both naturally and artificially synthesised, and are often used as catalysts or as intermediates in laboratory or industry activity.

Organic contaminants of water 2.2.12.1

OVERVIEW

Like chemical contaminants, organic contaminants can be found in water systems and pose a threat to the public and the environment.

THEORY DETAILS

In addition to chemical contamination of the waterways, there can also be organic contaminants found in the waterways. Organic contaminants that are found in the environment and resistant to being broken down are referred to as 'Persistent Organic Pollutants', or POPs. As these pollutants are resistant to biodegradation, these contaminants are sometimes referred to as 'forever chemicals'. Consequently, these molecules are able to accumulate in the environment leading to potential adverse effects on both human health and the environment. The following POPs were recognised as having detrimental effects on human health and the ecosystem, and grouped into three categories:

- **Pesticides**: aldrin, chlordane, DDT, dieldrin, endrin, heptachlor, hexachlorobenzene, mirex, toxaphene.
- **Industrial chemicals**: hexachlorobenzene, polychlorinated biphenyls (PCBs).
- **By-products**: hexachlorobenzene, polychlorinated dibenzo-p-dioxins and polychlorinated dibenzofurans (PCDD/PCDF), and PCBs.

As can be seen, the primary sources of POPs in the environment include pesticides in farming practice, chemicals from industrial run-off, and by-products from various sources such as domestic activities or from experiments conducted at schools, which are required to be disposed of with professional waste management.

In addition to their resistance to biodegradation and their persistent bioaccumulation, POPs are also of global concern due to their capability for long-distance transport. These POPs do not easily dissolve in water, and instead attach to molecules and settle in sediment. When aquatic animals consume these particles, they are difficult to break down, thus resulting in bioaccumulation down the food chain.

When persistent organic pollutants are consumed by humans, they remain in the body for extensive periods of time. This is because POPs are generally non-polar, and are therefore stored in the non-polar fatty tissue, rather than dissolving into the bloodstream and being filtered out via the kidneys. This presence of toxic molecules in the body for extensive periods of time can have severe detrimental effects on a person's health, including chronic illness and developmental defects.

Although there are now many restrictions and even bans on various forms of POPs, individuals can still have significant exposure due to either accidental production or as a result of an individual's occupation. However, for the majority of the population, most exposure to POPs is in the food that they consume due to bioaccumulation in fats of animals.

We will now cover the different categories of POPs in greater detail.

Pesticides

A pesticide is an agent that either kills or deters pests, and is used extensively in farming practices as shown in figure 8. There are various types of pesticides used in agricultural practice, each with different purposes. The targets of these different pesticides are as follows:

- Herbicide – Weeds and other plants
- Insecticides – Insects and arthropods
- Pheromones – Disrupting mating of insects



Image: Jinning Li/Shutterstock.com

Figure 8 Spraying pesticides.

Some of the pesticides used are POPs, able to be used due to their toxic properties and their effect on pests. In fact, the most commonly found POPs are organochlorine pesticides, such as DDT, which were used extensively from the 1940s through to the 1960s. However, the use of such pesticides ceased in the 1960s as they are a neurotoxicant, capable of causing short term effects, such as convulsions, headaches and vomiting, as well as long-term effects capable of damaging internal organs.

Dioxins

Dioxins are a family of chemical compounds that contain chlorine atoms and a particular carbon-oxygen structural framework. This is shown in figure 9.

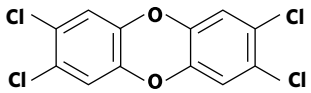


Figure 9 Dioxin.

These dioxins are primarily unwanted by-products produced in industrial processes, for example the smelting of metals, paper production, and the disposal of toxic waste. These dioxins can be highly toxic substances, although this toxicity has a large amount of variability. For the most part, dioxins are toxic and have a slow rate of breakdown, meaning that they remain in systems for extensive periods of time.

A large number of health effects of dioxins have been documented in medical literature, and they all place dioxins among the most toxic chemicals known. Dioxins are classified as known human carcinogens, but they are also capable of inducing non-cancerous effects such as high blood pressure and diabetes. Additionally, long-term exposure to dioxins can also cause disruption of the nervous, immune, reproductive, and endocrine systems.

In order to prevent the release of dioxins into the environment, toxic waste incinerators operate at higher than normal temperatures, causing further decomposition of toxic material in an attempt to break down the structures of dioxins before they can be released into the environment.

Oil spills

As discussed in lesson 6B, oil reserves are a source of a large quantity of energy rich hydrocarbons, which can be extracted for human usage. Because of this, when an oil reserve is located under the seabed, an oil rig is established above it to drill the oil for industrial purposes. In addition to this, extensive pipelines used to transfer the oil to refineries for collection are also constructed. However, these structures are not perfect, and in some cases the drilling or transport of oil causes oil spills as shown in figure 10, which is a form of organic contamination.



Image: Tigergallery/Shutterstock.com

Figure 10 Oil spill at Koh Samet.

As oil is non polar, when it spills it floats on the water forming a 'slick'. This has the benefit of preventing the spread of oil spills through the use of physical barriers called booms, protecting the shoreline and allowing some smaller constituents of the oil to evaporate. However, as there is a layer of oil on the water, this also has the effect of preventing sunlight from reaching the seabed, preventing many ecological processes from occurring underwater. Additionally, the oil will stick to anything it comes into contact with, including wildlife, shoreline rocks and sand, covering these with oil. This is particularly dangerous near shorelines, as the shallow waters often serve as nurseries for marine life. If marine life ingests the oil, which is considered poisonous, this can have devastating effects for local animal populations.

To prevent the formation of oil slicks, dispersants can be sprayed on oil to break it up into sludge, which eradicates the surface oil slicks. However, this use of dispersants can lead to further damage as oily sludge sinks to the bottom of the ocean. Instead, sometimes biological agents are used, such as added fertiliser sprayed on the oil, which fosters the growth of microorganisms that can feed on the oil and break it down.

Theory summary

- There are many different forms of water contamination.
- Water contamination can damage human health and the environment.
- Water sampling is carried out to assess for water contamination.
- Chemical contaminants such as salts can cause water hardness, heavy metal poisoning and can bioaccumulate.
- Organic contaminants includes pesticides, dioxins and oil spills.
- POPs are toxic organic contaminants that bioaccumulate.

10B QUESTIONS

Theory review questions

Question 1

Water quality sampling should be conducted

A once at the same depth for consistency.

B multiple times at different depths as the composition of water may vary.

Question 2

Water quality sampling is more likely to be conducted

A in the winter due to the higher solubility of contaminants in colder temperatures.

B following events that may introduce pollutants, such as bushfires.

Question 3

Run-off from agriculture and industry are primary sources of chemical contamination of water.

A True

B False

Question 4

Heavy metal poisoning is dangerous because

A heavy metals do not react with the body.

B heavy metals can disturb the structures of proteins essential for key bodily structures and processes.

Question 5

Organometallic compounds are

A molecules containing at least one carbon-metal bond.

B organic molecules showing intermolecular forces with metal ions.

Question 6

Hard water produces less lather from soap because

A metal cations neutralise stearate anions before they can react with dirt.

B heavier water requires more energy to transition to the gaseous state of foam.

Question 7

Bioaccumulation is the process whereby smaller animals die from heavy metal poisoning, enabling higher species to consume their remains.

A True

B False

Question 8

Persistent Organic Pollutants

A have been greatly restricted due to their high toxicity.

B are insignificant byproducts of industrial processes.

Question 9

Dioxins are

A toxic substances used as a fertiliser in agriculture.

B a toxic waste product from various industrial processes.

Question 10

Oil spills cause minimal damage as they are broken down by microbes and only float on the water surface.

A True

B False

Deconstructed questions

Use the following information to answer questions 11–13.

The Environment Protection Authority (EPA) carries out the essential function of protecting waterways and ensuring water quality is maintained in Australia, so that it may be consumed by the community without harm. Some important parameters that water scientists look for include the concentration of metal ions and evidence of undesired bacteria.

Question 11 (1 MARK)

Which of the options describes the water sampling method undertaken by the EPA?

A Collect a single sample from any point of a water source.

B Collect various samples at different locations and depths, changing these locations each time they sample the water source to assess a different part each time.

C Collect samples at various locations and depths of a water source, ensuring that the same materials and methods are used for each sample.

D Collect various samples, using different materials to store the water based on the turbidity of the water observed.

Question 12 (1 MARK)

What observations would indicate the presence of an algae bloom?

A Large quantities of nutrients such as nitrogen and phosphorous.

B Decreased levels of fish.

C Heightened presence of chlorophyll-A in the water producing a green colour.

D All of the above.

Question 13 (2 MARKS)

A water sample was found to produce reduced amounts of lather when soap was used with it. What is this type of water referred to as, and explain what the results of metal ion contamination and algae contamination tests would be expected to indicate.

Exam-style questions

Within lesson

Question 14 (12 MARKS)

Heavy metal contamination of water is a parameter strictly monitored by water authorities in Australia due to its potential to cause massive bodily harm.

a Provide one example of a heavy metal that may be found in a water supply. (1 MARK)

b It is known that when heavy metals are consumed, they can cause cancers, or organ and nervous system damage.

i How do heavy metals differ from other metals? (1 MARK)

ii Explain why heavy metals are dangerous when consumed, with reference to interactions with body structures. (3 MARKS)

c Provide two sources that may lead to heavy metals being found in water. (2 MARKS)

d Bioaccumulation is a phenomenon that is commonly observed in water contaminants.

i Describe the process of heavy metal bioaccumulation. (3 MARKS)

ii Draw a basic graph demonstrating heavy metal bioaccumulation. (2 MARKS)

Question 15 (10 MARKS)

Persistent Organic Pollutants (POPs), a series of organic pollutants and potential contaminants, have historically been used for a number of different uses, for example the use of DDT as a pesticide. However, the use of POPs was largely restricted in the Stockholm Convention on Persistent Organic Pollutants, as they can exhibit toxic effects on the population and the environment.

a POPs are sometimes referred to as 'forever chemicals'. Briefly explain why POPs have been given this name. (1 MARKS)

b Why does POPs' status as 'forever chemicals' make them particularly harmful? (2 MARKS)

c Give two examples of POPs. (2 MARKS)

d Explain the purposes of pesticides and why the use of POPs for this purpose can be dangerous. (2 MARKS)

e Explain how the chemical structure of POPs makes them remain in the human body for long periods of time. (3 MARKS)

Question 16 (8 MARKS)

The Deepwater Horizon disaster of 2010 was a large oil spill caused by an explosion on an oil rig off the Gulf of Mexico. The result of this was that vast amounts of oil were released into the water in the region.

a Is crude oil a polar or non-polar substance? (1 MARK)

b Describe how this oil spill would be expected to appear to observers. Explain your reasoning. (3 MARKS)

c Oil spills are a type of organic contamination. Explain why these oil slicks are dangerous for the surrounding environment. (2 MARKS)

d Scientists often spray biological agents onto oil spills to promote the growth of microorganisms capable of breaking down the oil spill. Discuss the efficacy of this method of removing oil spills. (2 MARKS)

Question 17 (8 MARKS)

The following article is taken from the Government of Western Australia Department of Health website:

> Safe drinking water supplies are critical to our health and wellbeing. Water from untreated sources such as streams, bores, dams and rainwater tanks can sometimes contain harmful microorganisms that cause serious illness or, in some cases, death.
>
> Chlorine has been used around the world for more than 100 years to disinfect drinking water supplies. As a result many water-borne diseases such as cholera, dysentery and typhoid have been effectively controlled.
>
> *Source: https://ww2 health.wa.gov.au/Articles/A_E/Chlorinated-drinking-water*

a Provide two examples of microorganisms that may contaminate drinking water. (2 MARKS)

b Is chlorine a contaminant in this case? Why, and can this change? (3 MARKS)

Peter, a chemistry student, believes that since chlorine is found in salt, the water containing intentionally added chlorine must now be saline.

c What would happen if someone were to consume saline water, such as seawater, without desalination? (2 MARKS)

Multiple lessons

Question 18 (11 MARKS)

Soaps and detergents use similar chemicals to remove dirt and other substances from materials.

a Describe how soaps remove dirt from materials, and the observation that can be made of this process. (3 MARKS)

b Using your knowledge of solubility and dissolution, explain why the dirt-removing chemicals are able to work in water and comment on their solubility. (3 MARKS)

c What is the name of water containing large amounts of metallic ions, and why does this water reduce the efficacy of soap? (3 MARKS)

d Why do some detergents come with anions added to the solution? (2 MARKS)

Question 19 (10 MARKS)

Water contaminants include a wide variety of different molecules, including both chemical and organic contaminants.

a Define the term 'organic molecule'. (1 MARK)

b Contaminants can enter water as water proceeds through the water cycle.

i What is the water cycle? (1 MARK)

ii What is the difference between freshwater, brackish water and seawater, and which of these is safe to consume? Explain. (3 MARKS)

c Various salts can be introduced to water through the passage of water through rocks and minerals, picking up ions during its passage.

i Can salts be considered a contaminant, even though they do not come from a human source? Justify your answer. (3 MARKS)

ii Are salts introduced to water from natural sources organic? Justify your answer. (2 MARKS)

Key science skills

Question 20 (10 MARKS)

A chemistry student hoping to one day become a water scientist sets up a series of experiments at home to perform analyses on a series of samples they collected from the Yarra River. The tests carried out by the student include electrical conductivity testing, as well as assessing turbidity and presence of microorganisms.

a The student collected three different samples at the same location, but different depths of the river. Is this a good representation of the water composition throughout the entire river? Explain. (3 MARKS)

b While the student collected each sample with the same technique, he unknowingly used a different container for one of the samples, which was reactive with the sample. What type of error is this? (1 MARK)

c Noticing their error, the student decides to collect three new samples at each depth, using identical containers. However, each of these containers had been used to store a single water sample and had not been washed properly, contaminating the sample to the same extent. What type of error is this? (1 MARK)

d How would the contamination in part c affect the accuracy and precision of the results collected respectively? (2 MARKS)

e List three water sampling protocol considerations that must be made to collect a legitimate water sample. (3 MARKS)

Exam-style question hints

14a: Heavy metals are metals not safe for human consumption. **14b:** The term heavy metals is indicative of their properties. **14bii:** Heavy metals have dangerous interactions with body proteins. **14c:** Many human processes utilise heavy metals. **14di:** Bio- refers to life, -accumulation refers to the build-up of a substance. **14dii:** x-axis is position on the food chain, y-axis is heavy metal concentration. **15a:** Forever as a reference to a long time. **15b:** The accumulation of dangerous substances can be harmful over time. **15c:** POPs can be either used or produced through human processes. **15d:** The use of POPs can lead to direct exposure or it can enter run-off. **15e:** Blood contains polar molecules. **16a:** Crude oil consists of hydrocarbons. **16b:** Water is a polar substance. **16c:** Oil spills are sticky and block sunlight. **16d:** Oil spills are very big, and microorganisms are very small. **17a:** Microorganisms may be naturally occurring in the water source or may be introduced. **17b:** Contaminants make something unfit for use. **17c:** The human body does not process salts effectively. **18a:** Soaps contain sodium stearate molecules. **18b:** Dissolution requires overcoming a solute and solvent's intermolecular bonds. **18c:** Metal ions are cations. **18d:** Stearate ions are anions. **19a:** Organic does not mean natural. **19bi:** Water is always in flux. **19bii:** Seawater is very salty. **19ci:** Contaminants make something unfit for use. **19cii:** Organic does not mean natural. **20a:** The composition of a water source can change between locations. **20b:** This was a one-off error. **20c:** This error affected all of the results to the same degree. **20d:** The error affected all of the results to the same degree. **20e:** Consistency is essential when collecting water samples.

10C ANALYSING WATER USING GRAVIMETRIC ANALYSIS

In this lesson, we will be learning about the use of gravimetric analysis and salinity testing for chemical analysis of a sample.

10A Water on Earth | 10B Water contamination | 10C Analysing water using gravimetric analysis | 10D Analysing water using light - part 1 | 10E Analysing water using light - part 2 | 10F Volumetric analysis | 10G Chromatography

Study design dot points

- sources of salts found in water (may include minerals, heavy metals, organo-metallic substances) and the use of electrical conductivity to determine the salinity of water samples
- the application of mass-mass stoichiometry to gravimetric analysis to determine the mass of a salt in a water sample

Key knowledge units

Determination of water salinity	2.2.8.2
Principles of gravimetric analysis	2.2.9.1
Calculations involving gravimetric analysis: mass to mass stoichiometry	2.2.9.2

Key terms and definitions

Gravimetric analysis chemical analysis involving the weighing of a precipitate formed from an analyte

Analyte substance which is quantified or identified

Reactant (in the context of precipitation) substance that reacts with an analyte to form a precipitate

Stoichiometry the relationship between relative quantities of reactants and products in a chemical reaction

Law of conservation of mass law that states that in a closed system, the total mass of reactants equals the total mass of products

Mass-mass stoichiometry calculation of reactants and products in a chemical reaction using the law of conservation of mass

Lesson links

This lesson builds on:

▶ 2F The mole - part 1

Moles are used in stoichiometric calculations.

▶ 7F Precipitation reactions

Gravimetric analysis involves precipitation reactions.

Determination of water salinity 2.2.8.2

OVERVIEW

Water analysis testing for salinity must be carried out to ensure appropriate usage.

THEORY DETAILS

As discussed in lesson 10A and 10B, only water containing very small concentrations of salt ions is safe for human consumption. Therefore, it is essential to carry out salinity testing.

Salinity testing involves determining the concentration of dissolved ions in a water sample, which is done by measuring the electrical conductivity of a sample of water. This is an effective means of water analysis as pure water is a poor conductor of electricity, while saline water with large quantities of dissolved ions will have high electrical conductivity. How well a water sample conducts electricity provides a good indication of the concentration of ions present, as samples which are more electrically conductive must have a higher concentration of ions.

The test to measure electrical conductivity requires two electrodes, a battery, a lightbulb and an ammeter, as shown in figure 1.

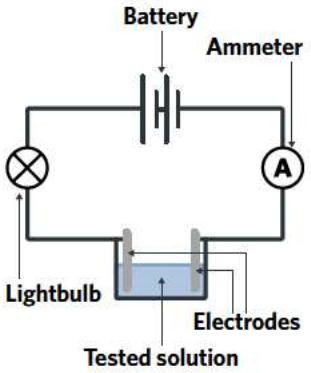


Figure 1 Solution electrical conductivity test.

As is visible in figure 1, electricity needs to be able to flow through the water from one electrode to the other in order to complete the circuit, and thus light up the lightbulb. The more electricity that is able to pass through the lightbulb, the brighter the light will appear. Therefore, the intensity of the light produced can be used as a measure of the salinity of the water. In order to obtain a more accurate measurement, an ammeter is typically used in the circuit to measure the amount of current passing through.

As is visible in figure 2, the conductivity of the water sample increases linearly as the concentration of ions increases, which can be measured using light intensity or ammeter readings. The unit of measurement for conductivity is microsiemens per centimetre (μS/cm).

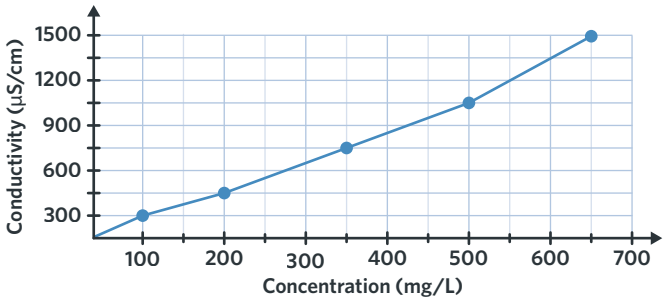


Figure 2 Concentration vs conductivity graph.

It should also be noted that the conductivity of a solution increases as the temperature increases, so it is standard practice to take electrical conductivity readings at 25°C to ensure that the results produced are consistent. In addition, conductivity readings simply provide an indication of the amount of ions found in the solution, and do not provide any qualitative information about the types of ions present.

Therefore, while not providing a comprehensive analysis of the types of ions found in a water sample, electrical conductivity testing can provide researchers with an indication of the salt concentration in a water sample. This enables researchers to determine whether a water source is safe for a particular usage.

Salinity meters come as handheld devices to assess water salinity outdoors as shown in figure 3. These are often used by the agricultural industry and water authorities to obtain a salinity estimate of a water source. However, as conditions are not as controlled outside of a lab, salinity meters must be calibrated for use.



Image: Sitthipong Pengjan/Shutterstock.com

Figure 3 Portable digital salinity meter.

Principles of gravimetric analysis 2.2.9.1

OVERVIEW

Gravimetric analysis is an analytical technique used to determine the quantity of a substance in a water sample.

THEORY DETAILS

Another method of determining the concentration of salt in a sample of water is through **gravimetric analysis**. In gravimetric analysis, salt ions are precipitated out of a solution to be weighed and to determine the quantity of the substance in a sample. The primary advantage of gravimetric analysis as opposed to electrical conductivity testing is that gravimetric analysis can be performed to specifically measure the amount of a certain ion in a solution, rather than measuring the total ion concentration in the water sample.

In gravimetric analysis, the substance whose quantity is being measured, the salts in the water sample, are referred to as the **analyte**. Depending on the identity of the analyte, a **reactant** is selected which will specifically react with the analyte. By weighing the precipitate formed by the analyte and the reactant, and determining the percentage of the mass of the precipitate made up by the analyte, the amount of analyte in the solution can be determined.

If other ions not being measured are also present in the water sample, the reactant must be chosen so that it doesn't form a precipitate with these ions, as this will lead to inaccuracies in measuring the amount of analyte in the water sample.

When adding reactant to the solution to form a precipitate, it is important to ensure that the reactant is added in excess. Since we are trying to determine the mass of the analyte in the water sample, we want to make sure that all of the analyte will be found in the precipitate. By adding the reactant in excess, this ensures that there will be sufficient amounts of reactant present for all of the analyte to react with, and the leftover reactant will remain dissolved in the solution without contributing to the mass of the precipitate.

Figure 4 is a summary of the method involved in gravimetric analysis.

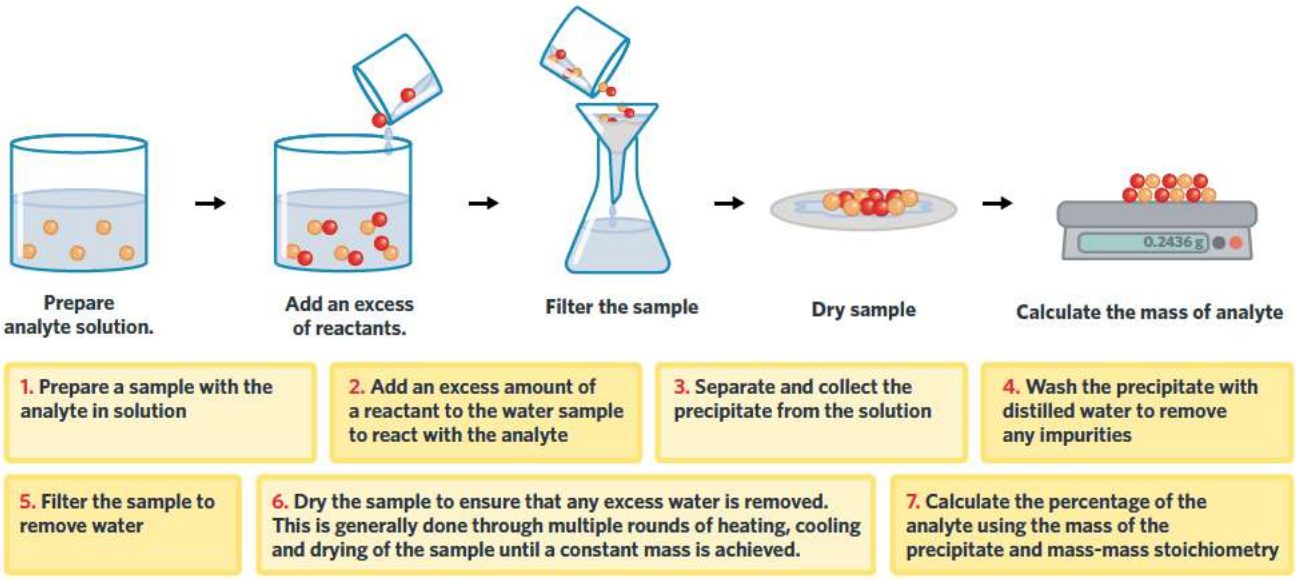


Figure 4 Gravimetric analysis technique.

Note that the precipitate should remain stable while being heated, as decomposition of the precipitate during heating will mean that the amount of analyte in the water sample cannot be accurately measured.

The formation of a precipitate from the analyte and reactant is important for isolating and measuring the amount of analyte in a substance, which may be either a solid or in a solution. If the analyte is found as a solid by itself, then the process of gravimetric analysis is unnecessary since the sample of pure analyte can be weighed by itself to determine how much of it is present. However, if the sample that the analyte is from contains multiple substances, then it will be impossible to determine the mass of a single analyte from the substance. Therefore, by dissolving all the contents of the sample in a solvent and selecting a reactant that only reacts with the analyte of interest, a precipitate containing only the analyte and the reactant will be formed, enabling quantitative analysis of the analyte to be performed. Lastly, if possessing a sample of analyte already dissolved in a solvent, it is important to precipitate the analyte out so that it can be weighed without the solvent.

Gravimetric analysis serves a number of important functions for scientists. As it is relatively inexpensive to find an ion to form a precipitate and measure the mass of the precipitate, gravimetric analysis can be carried out rapidly and easily. This has led to its introduction in the analysis of water impurities, foods, and other industrial uses.

Calculations involving gravimetric analysis: mass to mass stoichiometry 2.2.9.2

OVERVIEW

Mass-mass stoichiometry utilises knowledge of ratios of ions to determine the mass of an analyte in gravimetric analysis.

THEORY DETAILS

Having produced a precipitate from the sample solution, the next step is to use mass-mass stoichiometry to determine the amount of analyte that is in the precipitate. This is necessary since the precipitate consists of both the analyte and the reactant that was added, which cannot be individually weighed. In order to do so, we first need to revise our understanding of the concepts of moles.

As discussed in chapter 2, the mole is a unit of measurement for the amount of a substance. By knowing the mass and molar mass of a substance, the number of moles can be determined. Conversely, by knowing the number of moles of a substance, one can determine the mass of the substance present. These conversions between units can be carried out using the following formula:

$$\text{Number of moles} = \frac{\text{Mass (g)}}{\text{Molar mass (g per mol)}}$$

Another important concept in mass-mass stoichiometry is mole ratios. Take the equation below as an example:

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

The numbers in front of each molecule, the coefficients, indicate how many of that molecule will react each time the reaction occurs. In the example given, each time the reaction takes place, one $H_2SO_{4(aq)}$ molecule will react with two $NaOH_{(aq)}$ molecules. More importantly, the coefficients also indicate the mole ratios of reactants and products in a reaction. This concept of using mole ratios to determine quantitative data about any of the species involved in a reaction is referred to as **stoichiometry**.

For the reaction between $H_2SO_{4(aq)}$ and $NaOH_{(aq)}$, the mole ratio between the two reactants, $n(H_2SO_4) : n(NaOH)$ is 1 : 2. This illustrates that for every mole of $H_2SO_{4(aq)}$ used in the reaction, two moles of $NaOH_{(aq)}$ will be consumed. Mole ratios can be used to predict the amount of substances in reactions as long as the amount of one molecule is known, referred to as the 'known chemical'.

Another important concept to understand in stoichiometric calculations is that the total mass in the equation does not change, meaning that the total mass of the reactants will be equal to the total mass of the products. This is referred to as the **law of conservation of mass**. This is because no matter is being lost in the process of a reaction, it is simply being transferred to a different state.



When comparing substances, use the mole as the point of comparison.

Mass-mass stoichiometry

Mass-mass stoichiometry is a method of calculating the mass of reactants and products in a reaction using stoichiometry as described above. Let's apply this to gravimetric analysis by introducing a reaction that we may see in gravimetry, as shown:

$Pb(NO_3)_{2(aq)} + 2KCl_{(aq)} \rightarrow PbCl_{2(s)} + 2KNO_{3(aq)}$

As we can see, the reaction between the two solutions lead nitrate ($Pb(NO_3)_{2(aq)}$) and potassium chloride ($KCl_{(aq)}$) results in the formation of a lead chloride precipitate ($PbCl_{2(s)}$). If we compare the mole ratios between the reactants and precipitate, we will identify different relationships as seen in table 1.

Table 1 Breakdown of the chemical equations between lead nitrate and potassium chloride.

Reference substance	Substance being compared to	Mole ratio
$Pb(NO_3)_{2(aq)}$	$PbCl_{2(s)}$	$n(Pb(NO_3)_2 : n(PbCl_2)$ 1 : 1
$KCl_{(aq)}$	$PbCl_{2(s)}$	$n(KCl) : n(PbCl_2)$ 2 : 1

We can use this information to perform appropriate calculations from our reaction involving lead nitrate $Pb(NO_3)_{2(aq)}$ and $KCl_{(aq)}$. Let's assume that 127.0 g of precipitate was formed. If we were required to calculate the mass of $KCl_{(aq)}$ needed to produce this mass of $PbCl_{2(s)}$ precipitate, we would use the steps shown in figure 5.

1	Calculate the amount of the known substance.	1	$n(PbCl_2) = \frac{m}{M}$ $n(PbCl_2) = \frac{127.0}{207.2 + (35.5 \times 2)} = 0.457$ mol
2	Use the balanced equation to find the mole ratio between the known and unknown substance.	2	$n(PbCl_2) : n(KCl)$ 1 : 2 0.457 : ?
3	Calculate the amount of the unknown.	3	$n(KCl) = \frac{2}{1} \times 0.457$ mol = 0.914 mol
4	Calculate the mass of the unknown.	4	$m(KCl) = n \times M$ $m(KCl) = 0.914 \times (39.1 + 35.5) = 68.2$ g

Figure 5 Application of mass to mass stoichiometry to calculate an unknown mass.

Worked example 1

A reaction occurred between MgCl and Na_2CO_3 forming a precipitate as seen below:

$MgCl_{2(aq)} + Na_2CO_{3(aq)} \rightarrow MgCO_{3(s)} + 2NaCl_{(aq)}$

If 21.1 g of $MgCl_2$ reacted with an excess of Na_2CO_3, determine the mass of $MgCO_3$ that would be formed.

What information is presented in the question?

The mass of $MgCl_2$ and the identity of the reactants and products.

What is the question asking us to do?

Determine the mass of $MgCO_{3(s)}$ precipitate formed.

What strategy(ies) do we need in order to answer the question?

1. Determine the amount of $MgCl_2$ reacted.
2. Compare the mol ratio between $MgCl_2$ and $MgCO_3$.
3. Determine the number of moles of $MgCO_3$ that would be formed.
4. Determine the mass of $MgCO_3$ using the molar mass and number of moles.

Answer

$m(MgCl_2) = 21.1$ g

$M(MgCl_2) = 24.3 + (35.5 \times 2) = 95.3$ g mol^{-1}

$n(MgCl_2) = \frac{21.1}{95.3} = 0.221$ mol

$n(MgCl_2) : n(MgCO_3) = 1 : 1$

$n(MgCO_3) = 0.221$ mol

$M(MgCO_3) = 24.3 + 12.0 + (3 \times 16.0) = 84.3$ g mol^{-1}

$m = n \times M = 0.221 \times 84.3 = 18.7$ g

By using the amount of a known substance, we can use stoichiometry to determine the amount of a different substance in the reaction, and use that information to find the mass of the substance.

Mass-mass stoichiometry in gravimetric analysis

We can also use our understanding of the general concepts of mass-mass stoichiometry to find the composition or concentrations of substances.

Continuing with our example shown in figure 5, suppose we were told that the $PbCl_{2(s)}$ precipitate was collected from an unknown substance that had a total mass of 219.6 g. We can calculate the percentage mass of chlorine found in the original sample using the steps shown in figure 6.

5	Determine the molar ratio between the known substance and the substance we're looking for.	5	$n(KCl) : n(Cl)$ $1 : 1$
6	Determine the amount, in mol, of the substance we're looking for.	6	$n(KCl) = n(Cl)$ $n(Cl) = 0.914$ mol (as calculated above)
7	Determine the mass of the substance we're looking for.	7	$m(Cl) = 0.914 \times 35.5 = 32.4$ g
8	Calculate the percentage.	8	$\%(Cl) = \frac{32.4}{219.6} \times 100 = 14.8\%$

Figure 6 Determining the percentage of a substance in a sample using mass-mass stoichiometry.

The assumption in this scenario is that all the chlorine atoms present existed in the form of $KCl_{(aq)}$. This will generally be the case unless we are told otherwise.

By using this method, we can quantitatively determine the presence of a particular substance in a sample.

Worked example 2

A 3.00 g sample of limestone ($CaCO_3$) was dissolved, where the calcium ions were precipitated using excess ammonium oxalate solution $(NH_4)_2C_2O_{4(aq)}$. The precipitate formed was calcium oxalate, $CaC_2O_{4(s)}$ according to the equation:

$Ca^{2+}_{(aq)} + C_2O_4{}^{2-}_{(aq)} \rightarrow CaC_2O_{4(s)}$

Determine the percentage, by mass, of calcium present in limestone if 3.65 g of precipitate was formed.

What information is presented in the question?

The precipitation reaction, the mass of limestone (3.00 g) and the mass of precipitate formed (3.65 g).

What is the question asking us to do?

Determine the percentage, by mass, of calcium present in limestone.

What strategy(ies) do we need in order to answer the question?

1. Calculate the amount of precipitate formed.
2. Compare the molar ratio between the precipitate and calcium.
3. Determine the amount of calcium present.
4. Calculate the mass of calcium present.
5. Calculate the percentage by mass of calcium in the limestone sample.

Answer

$$n(CaC_2O_4) = \frac{3.65}{40.1 + (2 \times 12.0) + (4 \times 16.0)}$$

$= 0.02849$ mol

$n(CaC_2O_4) : n(Ca^{2+})$

$1 : 1$

$n(Ca^{2+}) = 0.02849$ mol

$m(Ca^{2+}) = 0.02849 \times 40.1 = 1.14$ g

$$\%Ca^{2+} = \frac{1.14}{3.00} \times 100 = 38.1\%$$

Therefore, the limestone sample contains 38.1% calcium.

Mass-mass stoichiometry involving excess reactant

With a deeper level of understanding of chemical reactions, we now know that chemical reactions occur as per a given ratio. Therefore, not all of the substance that we add into a reaction will react. Sometimes, we have more of one reactant than the other, resulting in an **excess reactant**, and a **limiting reactant**. The limiting reactant is completely consumed in the reaction and is the limiting factor to how much product we can make whereas the excess reactant will be left over after the reaction is considered to be complete.

Imagine making a cheese sandwich consisting of two slices of bread and a single slice of cheese. Suppose we were given seven slices of bread and three slices of cheese. We can figure out the number of cheese sandwiches that can be made as shown in figure 7.

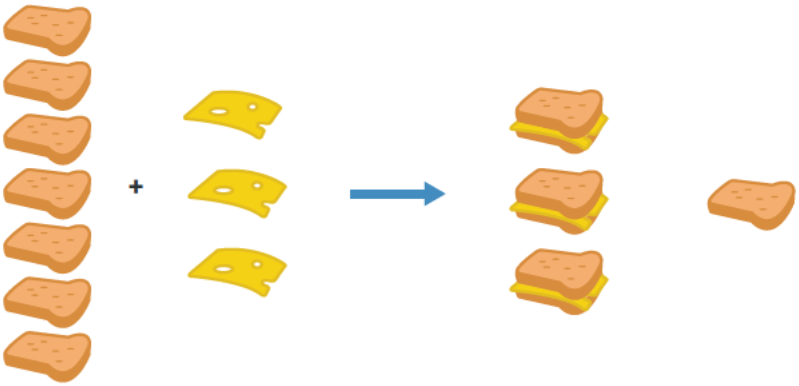


Figure 7 Illustration of excess and limiting reactants in a chemical reaction.

As we can see, we can make a total of three cheese sandwiches, where all of the cheese slices have been used leaving a leftover slice of bread. The slices of bread are considered to be our excess reactant, whereas the cheese slices are considered to be our limiting reactant. It's important to know that the products formed in a chemical reaction will depend on the amount of limiting reactant available.

In the context of chemical reactions, consider the reaction between methane and oxygen:

$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$

Suppose we were given three mol of $CH_{4(g)}$ and three mol of $O_{2(g)}$ for the reaction to occur as shown in figure 8.

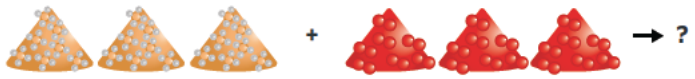


Figure 8 Sample of $CH_{4(g)}$ and $O_{2(g)}$ available for a reaction.

Using the molar ratio highlighted in the balanced equation, we can determine both the limiting and excess reactant. A method by which we can do this is to look at it from the perspective of one of the reactants. For this example, let's use $CH_{4(g)}$ as our reference point.

According to our balanced chemical equation, the molar ratio between $CH_{4(g)}$ and $O_{2(g)}$ is 1 : 2, where for every one mole of $CH_{4(g)}$ used, two moles of $O_{2(g)}$ are required. Given that we have three moles of $CH_{4(g)}$ we can determine how many moles of $O_{2(g)}$ are required as shown in figure 9.

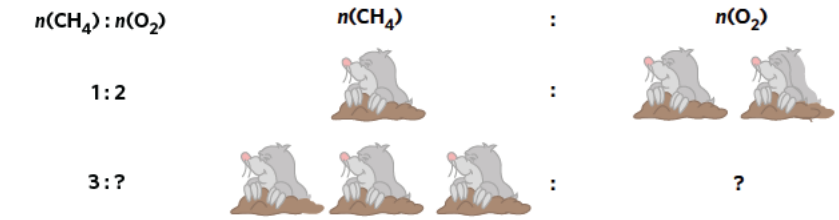


Figure 9 Determining limiting and excess reactants.

As we can see in figure 9, in order for all three moles of $CH_{4(g)}$ to be used, six moles of $O_{2(g)}$ is required. However we have only been given three moles of $O_{2(g)}$. We therefore do not have enough $O_{2(g)}$ to use all of the $CH_{4(g)}$ given. As such, $O_{2(g)}$ is our limiting reactant and will be completely consumed, whereas $CH_{4(g)}$ is the excess reactant. When finding the limiting and excess reactants, we must compare the amount, in mol, of substances.



To find the limiting and excess reactant, we need to compare the amount, in mol, of both reactants.



The excess reactant is always the reactant that is 'left over'.

Worked example 3

A student is carrying out a reaction between sulfuric acid and sodium hydroxide, which has the following equation:

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

If there is 90.0 g of H_2SO_4 and 70.0 g of NaOH present in the sample, determine which is the limiting reactant and which is the excess reactant.

What information is presented in the question?

The chemical equation and the masses of reactants.

What is the question asking us to do?

Determine the limiting and excess reactant in the reaction.

What strategy(ies) do we need in order to answer the question?

1. Determine the mole ratio of the reactants.
2. Determine the number of moles of each reactant.
3. Compare the number of moles using the mole ratio.
4. Determine which reactant is in excess and which is limiting.

Answer

$n(H_2SO_4) : n(NaOH)$

1 : 2

$$n(H_2SO_4) = \frac{90.0}{(2 \times 1.0) + 32.1 + (4 \times 16.0)} = 0.917 \text{ mol}$$

$$n(NaOH) = \frac{70.0}{23.0 + 16.0 + 1.0} = 1.75 \text{ mol}$$

If we use H_2SO_4 as a reference point:

$n(H_2SO_4) : n(NaOH)$

1 : 2

0.917 mol : x

$$n(NaOH) = \frac{2}{1} \times 0.917 = 1.83 \text{ mol}$$

To use all 0.917 mol of H_2SO_4, a total of 1.83 mol of NaOH is required. Considering there is only 1.75 mol of NaOH available, NaOH is the limiting reagent and H_2SO_4 is the excess reagent.

When there is a limiting reactant, we must use the limiting reactant as our point of reference to calculate the amount of all other substances in the reaction.

Worked example 4

A 5.23 g sample of $AgNO_3$ and 1.62 g sample of NaCl was used in the following reaction:

$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow NaNO_{3(aq)} + AgCl_{(s)}$

Determine the mass of AgCl precipitate that would be formed.

What information is presented in the question?

The chemical equation and the masses of reactants.

What is the question asking us to do?

Determine the mass of precipitate that would be formed.

What strategy(ies) do we need in order to answer the question?

1. Determine the amount of each reactant present.
2. Determine the limiting reactant.
3. Apply stoichiometry by using the limiting reactant to find the mass of precipitate formed.

Answer

$$n(AgNO_3) = \frac{5.23}{107.9 + 14.0 + (3 \times 16.0)} = 0.03078 \text{ mol}$$

$$n(NaCl) = \frac{1.62}{23.0 + 35.5} = 0.02769 \text{ mol}$$

If we use $AgNO_3$ as a reference point to determine the limiting reactant:

$n(AgNO_3) : n(NaCl)$

$1 : 1$

0.03078 mol : ?

$$n(NaCl) = \frac{1}{1} \times 0.03078 = 0.03078 \text{ mol}$$

Therefore to use 0.03078 mol of $AgNO_3$, we would require 0.03078 mol of NaCl. However, we only have a total of 0.02769 mol of NaCl, which is not enough. NaCl is our limiting reactant and will be our point of reference to calculate the mass of precipitate formed.

$n(NaCl) : n(AgCl)$

$1 : 1$

0.02769 mol : ?

$$n(AgCl) = \frac{1}{1} \times 0.02769 = 0.02769 \text{ mol}$$

$m(AgCl) = 0.02769 \times (107.9 + 35.5) = 3.97$ g

In this reaction, 3.97 g of AgCl precipitate will be formed.

Theory summary

- Salinity testing uses electrical conductivity principles of pure water and ions to determine the salinity of a water sample.
- Gravimetric analysis involves the formation and weighing of a precipitate to determine the amount of an analyte.
- Mass-mass stoichiometry uses the stoichiometric ratios in an equation to determine the amount/mass of other substances.
- Limiting reactants will completely react and determine the extent of a reaction, while excess reactants are left over at the end of the reaction.
- When calculating amounts of substance, always use the limiting reactant as the point of reference.

10C QUESTIONS

Theory review questions

Question 1

Electrical conductivity testing provides information about

A the concentration of ions in water.

B the amount of precipitate in water.

Question 2

Water samples with a high salt concentration have

A higher electrical conductivity.

B lower electrical conductivity.

Question 3

Lightbulbs receiving more electricity will

A shine for a longer duration of time.

B have a higher light intensity.

Question 4

What is gravimetric analysis?

A Chemical analysis technique involving interactions of ions with light.

B Chemical analysis technique involving precipitation of ions out of solution to determine their quantity.

Question 5

In gravimetric analysis, a poorly dried precipitate will result in a higher mass.

A True

B False

Question 6

A precipitate is formed in gravimetric analysis through

A the introduction of a reactant to an analyte in solution.

B supersaturating a solution to form crystals.

Question 7

Does gravimetric analysis provide quantitative or qualitative information?

A Qualitative

B Quantitative

Question 8

What is the molar ratio of NH_3:O_2 in the following reaction: $4NH_{3(g)} + 3O_{2(g)} \rightarrow 2N_{2(g)} + 6H_2O_{(l)}$

A 2 : 6

B 4 : 3

Question 9

If five mol of H_2 reacts with three mol of O_2 in the following reaction:

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

Which is the limiting reactant?

A H_2

B O_2

Question 10

If the molar ratio of ion A : ion B is 2 : 1, then the mass of chemical A is

A double the mass of ion B.

B unable to be determined with the information provided.

Deconstructed questions

Use the following information to answer questions 11–13.

Joseph is a chemistry professor assessing the chemical composition of a sample of seawater. He decides to use either electrical conductivity testing or gravimetric analysis as a means of measuring the concentration of sodium in sample of seawater.

Question 11 (1 MARK)

Which option correctly describes the method used for gravimetric analysis?

A Measuring the amount of electricity flowing through a circuit connected to a water sample.

B Inserting a sample into a flame and observing the colour produced to determine the chemical composition.

C Determining the amount of a substance dissolved in water through weighing a precipitate formed with it.

D Weighing a sample of water and subtracting the mass of water to determine the concentration of salts.

Question 12 (1 MARK)

Which option correctly describes the method used for electrical conductivity testing?

A Measuring the amount of electricity flowing through a circuit connected to a water sample.

B Inserting a sample into a flame and observing the colour produced to determine the chemical composition.

C Determining the amount of a substance dissolved in water through weighing a precipitate formed with it.

D Weighing a sample of water and subtracting the mass of water to determine the concentration of salts.

Question 13 (4 MARKS)

Suggest some advantages and disadvantages of using gravimetric analysis and electrical conductivity testing. Justify your reasoning.

Exam-style questions

Within lesson

Question 14 (16 MARKS)

Sometimes, not all reactants are used up in a chemical reaction.

a For each of the following reactions, determine the limiting reactant.

i $MgSO_{4(aq)} + Ca(NO_3)_{2(aq)} \rightarrow CaSO_{4(s)} + Mg(NO_3)_{2(aq)}$ with 0.332 mol of $MgSO_{4(aq)}$ and 1.98 mol of $Ca(NO_3)_{2(aq)}$. (2 MARKS)

ii $2AgNO_{3(aq)} + CaCl_{2(aq)} \rightarrow Ca(NO_3)_{2(aq)} + 2AgCl_{(s)}$ with 3.71 mol of $AgNO_{3(aq)}$ and 1.94 mol of $CaCl_{2(aq)}$. (2 MARKS)

iii $Pb(NO_3)_{2(aq)} + 2NaCl_{(aq)} \rightarrow PbCl_{2(s)} + 2NaNO_{3(aq)}$ with 0.866 g of $Pb(NO_3)_{2(aq)}$ and 2.34 g of $NaCl_{(aq)}$. (2 MARKS)

iv $4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$ with 1.45 g of $Al_{(s)}$ and 3.11 g of $O_{2(g)}$. (2 MARKS)

b For each of the reactions above, calculate the mass of precipitate that would be formed. (8 MARKS)

Multiple lessons

Question 15 (6 MARKS)

A scientist wants to determine the mass of aluminium nitrate in a solution by adding sodium phosphate.
The equation for this reaction is:

$Al(NO_3)_{3(aq)} + Na_3PO_{4(aq)} \rightarrow AlPO_{4(s)} + 3NaNO_{3(aq)}$

In his experiment, the scientist finds that 7.78 g of precipitate is obtained.

a What is the scientific name of the precipitate formed? (1 MARK)

b Calculate the number of moles of precipitate formed. (2 MARKS)

c Determine the molar ratio of the reactants in the reaction. (1 MARK)

d Calculate the mass of $Al(NO_3)_3$ in the solution. (2 MARKS)

Question 16 (9 MARKS)

In the chemistry lab at his school, Alan is attempting to determine the concentration of chloride ions in a sample of seawater. In the lab with him Alan has a standard solution of silver nitrate. When the silver nitrate is mixed with the seawater in excess, a white precipitate of silver chloride (AgCl) is formed.

a Why must the silver nitrate be added in excess? Explain your reasoning. (2 MARKS)

b What must be done before the precipitate can be weighed? Why? (3 MARKS)

c Alan removes the precipitate from the solution, following correct experimental technique, and records the mass to be 167.4 g.

i What is the mass of chloride in the sample of seawater? Show your calculations. (Assume all chloride ions were precipitated). (3 MARKS)

ii The sample of seawater contains two litres of water. What is the concentration of chloride ions in the seawater? (1 MARK)

Question 17 (14 MARKS)

Hard water is water that has a high concentration of cations present, which prevents soaps from forming a lather when introduced to water. Because of this, the efficacy of soaps in hard water is greatly reduced.

a Explain why soap is less effective in hard water, and why a precipitate is formed. (3 MARKS)

b A scientist suggests that in order to determine the concentration of ions that are present in a sample of hard water, the precipitate could be collected and weighed.

i Assuming that the scientist did not know the identity of the ions in the precipitate, would this be an effective method? Explain your answer. (2 MARKS)

ii Assuming that the scientist did know the identity of the ions in the precipitate, would this be an effective method for measuring the number of soap ions in the sample? Explain your answer. (3 MARKS)

iii The scientist carries out the experiment correctly . Determine whether the analyte or the added reactant would be the limiting reactant. (1 MARK)

c A 500 mL solution containing stearate anions ($C_{18}H_{35}O_2^-$) was produced from adding 300.0 g of soap. Magnesium cations (Mg^{2+}) were added using correct technique, leading to the formation of a solid with a mass of 61.0 g. The equation of the reaction taking place is:

$Mg^{2+}_{(aq)} + 2C_{18}H_{35}O_2{}^-_{(aq)} \rightarrow Mg(C_{18}H_{35}O_2)_{2(s)}$

i Determine the concentration of stearate anions in the solution. (3 MARKS)

ii Determine the percentage soap by mass of stearate anions in the soap. (2 MARKS)

Question 18 (8 MARKS)

A chemistry teacher is setting up an experiment for her class to practice stoichiometric calculations. She begins by adding 150.0 g of gypsum ($CaSO_4 \cdot 2H_2O$) to a beaker containing water, with all of the gypsum dissolving. To determine the amount of calcium in the solution, phosphoric acid (H_3PO_4) is added to the solution to carry out the following reaction:

$3Ca^{2+}_{(aq)} + 2PO_4{}^{3-}_{(aq)} \rightarrow Ca_3(PO_4)_{2(s)}$

Note: $CaSO_4 \cdot 2H_2O$ is a hydrate, which means it is a substance that contains water.

a If this experiment is carried out correctly, which reactant should be the limiting reactant? Explain your answer. (2 MARKS)

b After allowing the reaction to take place fully, the precipitate was removed from the solution.

i Identify one factor that must be accounted for before weighing the sample. (1 MARK)

ii The class measured the mass of the precipitate to be 87.4 g. Identify the number of moles of calcium that have reacted. (3 MARKS)

iii Calculate the %(m/m) of calcium in gypsum. (2 MARKS)

Question 19 (17 MARKS)

Vinh is a Year 11 Chemistry student who wants to put together his knowledge of chemical analytical techniques with other principles of chemistry. As a result, Vinh decides to set up the following experiment:

1 Add a sample of pure octane to water.

2 Add a reactant to form a precipitate with octane.

3 Dry the precipitate with a paper towel and weigh it.

4 Calculate the amount of octane through mass-mass stoichiometry.

a What term is used to refer to the substance that is being quantitatively analysed in gravimetric analysis? (1 MARK)

b Will it be possible to use octane in the experiment described? Justify your answer. (3 MARKS)

c Identify a potential change to the solvent that could be made to overcome the issue identified in Q19b. Justify your answer. (2 MARKS)

d Is it necessary to carry out the experiment at all? Explain your answer. (3 MARKS)

e Vinh decides to maintain the experimental technique outlined above, but replaces the analyte with a solution containing $MgCl_2$. Through reacting the solution with $AgNO_3$, a precipitate of AgCl was formed.

i With reference to the experimental technique outlined above, identify at least two potential sources of error. (2 MARKS)

ii Write the equation of the reaction between $MgCl_2$ and $AgNO_3$. (2 MARKS)

iii The solution contains 33.2 g of $MgCl_2$, and 108.0 g of $AgNO_3$ is added to form a precipitate. Using calculations, identify the limiting and excess reactant in this reaction. (3 MARKS)

iv Calculate the mass of precipitate that will be formed. (1 MARK)

Key science skills

Question 20 (5 MARKS)

In chemical analysis, it is vitally important to ensure that all collected values are accurate and precise in order to gather correct information about the sample being analysed.

a Define the terms accuracy and precision. (2 MARKS)

b A student conducting gravimetric analysis of a solution of copper ions forms a precipitate and weighs it. The student repeats this same method across a number of trials.

i Looking at his results the student notices that all of the measurments are close to each other but not close to the expected value. Comment on the precision and accuracy of these results. (1 MARK)

ii What type of error has the student most likely made? Justify your answer. (2 MARKS)

Exam-style question hints

14a: Compare the relative amounts of reactants to mole ratios. **14b:** The limiting reactant determines the quantity of products.**15a:** Ionic compounds are named based on the ions present. **15b:** Moles equals mass divided by molar mass. **15c:** Molar ratio is illustrated in the chemical equation. **15d:** The mass formed depends on the number of moles and the molar mass. **16a:** The reactant is not the substance being measured. **16b:** Sample contamination should be avoided where possible. **16c:** Molar masses are important for determining how much of a precipitate is made up by an element. **17a:** This molecule is often found in a compound with a metal ion. **17b:** Soap ions work best as ions. **17c:** Soap ions are the analyte, and salt ions are the reactant. **18a:** Consider which reactant is being measured. **18b:** A precipitate should not be contaminated when being weighed. **19a:** The substance under anaylsis has a specific name. **19b:** Polarity affects the ability of substances to interact. **19c:** Octane is a non-polar molecule. **19d:** Gravimetric analysis precipitates a substance out of a solution to be weighed. **19e:** All of the substance must be weighed for gravimetric analysis to be effective. **20a:** These variables measure different aspects of data. **20b:** The type of error depends on consistency of results.

10D ANALYSING WATER USING LIGHT - PART 1

In this lesson, we will be learning about colorimetry and UV-visible spectroscopy.

10A Water on Earth | 10B Water contamination | 10C Analysing water using gravimetric analysis | 10D Analysing water using light - part 1 | 10E Analysing water using light - part 2 | 10F Volumetric analysis | 10G Chromatography

Study design dot point

- the application of colorimetry and/or UV-visible spectroscopy, including the use of a calibration curve, to determine the concentration of coloured species (ions or complexes) in a water sample

Key knowledge units

Analysis of water using colorimetry	2.2.10.1
Analysis of water using UV-visible spectroscopy	2.2.10.2

Key terms and definitions

Colorimetry method of chemical analysis exploiting the light absorbance properties of substances with different concentrations

Colorimeter instrument used to undertake colorimetric analyses

Electromagnetic radiation waves of the electromagnetic field including visible and ultraviolet light

Spectroscopy analysis of the interaction between matter and electromagnetic radiation

Complementary colours colours opposite to each other on the colour wheel

Standard solution solution with a precisely known concentration

Calibration curve graph depicting the relationship between concentration and light absorbance for a particular substance

UV-visible spectroscopy spectroscopy method utilising the ultraviolet light region of the electromagnetic spectrum

UV-visible spectrophotometer instrument used in UV-visible spectroscopy

Monochromator device used to select a particular wavelength of light

Lesson links

This lesson builds on:

- 2A The atom

Differing atomic structures of elements cause different interactions with light.

Analysis of water using colorimetry 2.2.10.1

OVERVIEW

Colorimetry is an analytical technique that uses visible light to determine the quality of a sample of water.

THEORY DETAILS

As we have seen in previous lessons, water samples can have varying degrees of quality depending on the concentration of contaminants such as salts. In 10C we explored how electrical conductivity and gravimetric analysis can be employed to determine the concentration of ions in solution. We will be covering some water quality analysis techniques in this lesson that make use of the interactions between ions and light.

In this section, we will be introducing another method for chemical analysis known as **colorimetry**. This method is used as it is fast and non-destructive. The non-destructive nature of the technique means that the sample of water remains undamaged throughout the experiment.

Colorimetry involves measuring the intensity of colour in a solution, which is directly proportional to the concentration of ions in the solution. While the colour of a solution may be observable by eye in some cases, a device known as a **colorimeter** measures the concentration of ions in a solution more accurately. In order to understand the function of a colorimeter, we first need to investigate some fundamental principles of light.

Light

Visible light is a form of **electromagnetic radiation**. Other forms of electromagnetic radiation include x-rays, ultraviolet light and radio waves. The different types of light radiation are shown in figure 1.

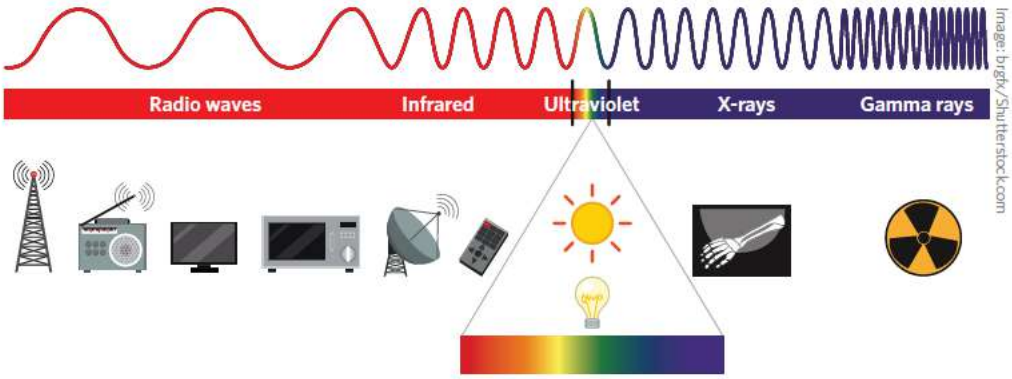


Image: brgfx/Shutterstock.com

Figure 1 Electromagnetic spectrum.

Electromagnetic radiation such as visible light interacts with atoms. By distinguishing between different types of atomic interactions with light, an experimenter is able to ascertain the atomic make-up of the solution being tested. This form of chemical analysis is referred to as **spectroscopy**. Colorimetry applies the principles of spectroscopy within the scope of visible light, which corresponds to wavelengths between 400 nm and 700 nm on the spectrum.

Materials are able to absorb and reflect different wavelengths of light depending on their chemical composition. The colour reflected by the substance is that which is visible to the observer. For example, a copper solution will absorb all colours except blue, which will be reflected into the observer's eyes. Hence, the solution will appear blue. This principle is demonstrated in figure 2, using a ladybug as an example.

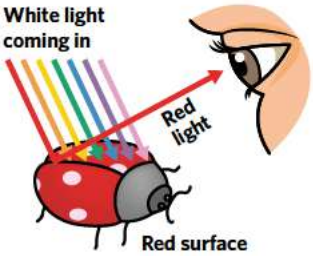


Figure 2 Reflected light.

The observed colour of a material and the colour most strongly absorbed are referred to as **complementary colours** – these are found on opposite sides of the colour wheel as shown in figure 3. For example, orange is a complementary colour to blue. This means that a blue solution will absorb orange light most strongly.



Figure 3 Colour wheel.

One method of learning the complementary colours is by remembering the order in which colours appear on the electromagnetic spectrum from the largest wavelength to the smallest wavelength. ROYGBV is an acronym commonly used to remember complementary colours which stands for red, orange, yellow, green, blue and violet. When each complementary colour is written below ROYGBV, the acronym GBVROY is produced. The pairs of complementary colours produced from the two acronyms can be lined up as shown in figure 4. As can be seen here, red and green, orange and blue, and yellow and violet are all sets of complementary colours.

ROYGBV
GBVROY

Figure 4 Alignment of complementary colours.

If the substance being tested is opaque (cloudy), the substance will reflect the complementary colour of the colour most strongly absorbed. The colour that is reflected is what we see. However, if the substance is transparent, the substance will freely transmit the colour complementary to that most strongly absorbed. The colour being transmitted is what we are able to see. These different interactions of light with a material are shown in figure 5.

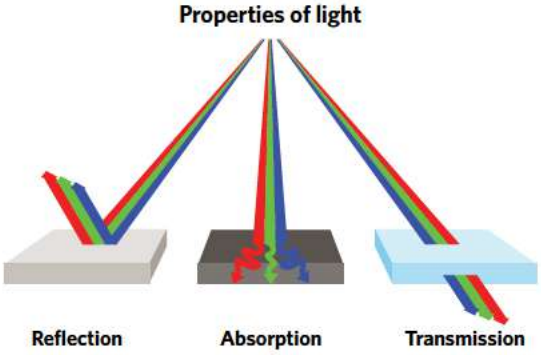


Image: OSweetNature/Shutterstock.com

Figure 5 Reflection, absorption and transmission of light.

Colorimetry

A substance's ability to reflect, absorb or transmit light can be examined to determine the concentration of ions in solution. This is the premise of colorimetry which uses a colorimeter to carry out this function.

Colorimeters function by measuring the absorbance of visible light as the light passes through a solution. This enables the concentration of solute in solution to be established. Each solute has its own unique atomic structure which means that every solution absorbs light best at a different wavelength. Furthermore, more concentrated solutions absorb more electromagnetic radiation. This is known as the Beer-Lambert Law which states that absorbance of light is proportional to the concentration of a substance. Hence, the more light that is absorbed, the greater the concentration of the sample.

Colorimetry is both a qualitative and quantitative technique. It is used qualitatively to determine the identity of a solute in a solution and quantitatively to determine the concentration of a solute in solution.

A colorimeter consists of three main parts used to measure the intensity of colour in a sample solution as shown in figure 6.

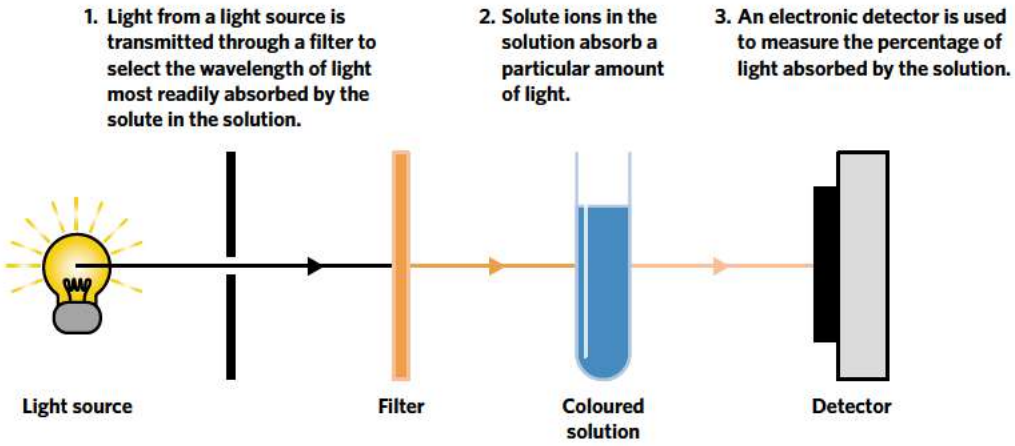


Figure 6 Analysis of solutions using colorimetry.

First, the light produced is filtered in order to select for the wavelength that will be most strongly absorbed by the sample. For the blue copper solution in figure 6, the filter allows only the complementary colour, orange light, to pass through. A recorder then detects how much light passes through the solution to generate a value which represents how much light was absorbed by the solution. We know that the greater the concentration of copper ions in the solution, the greater the absorbance of the orange light. This will lead to a higher percentage of light absorbed by the solution.

Calibration curves

To accurately determine the concentration of a solution using colorimetry, the absorbance results from the electronic detector must be compared to a series of **standard solutions**. Standard solutions contain precisely known concentration of solute ions. Developing a calibration curve from standard solutions allows us to identify the concentration of the sample solution. That is, the absorbance data from the standard solutions can be compared with the sample under analysis to determine the concentration of solute in the sample.

A **calibration curve** is used to visualise the relationship between concentration and absorbance. These calibration curves are a plot of concentrations of the solutions on the x-axis, and their corresponding absorbance values on the y-axis. This is shown in figure 7 which depicts a linear relationship between concentration and absorbance in accordance with the Beer-Lambert Law.

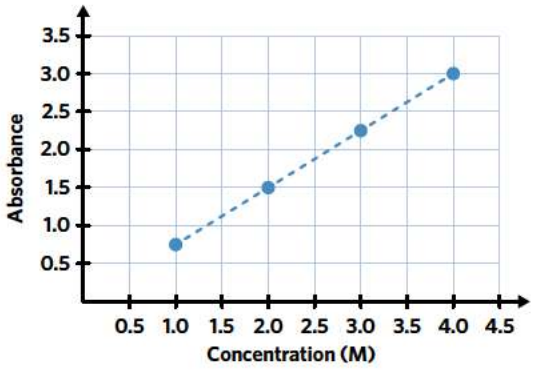


Figure 7 Calibration curve.

By constructing a calibration curve, we are able to ascertain the concentration of a sample solution using the absorbance value produced by the detector. To determine the concentration, the x-value (concentration) which corresponds to the y-value (absorbance) from the detector needs to be identified. For example, if we take a solution with a recorded absorbance of 2.5, we can then find its location on the curve as shown in figure 8. As can be seen, the concentration corresponding to an absorbance of 2.5 is approximately 3.35 M.

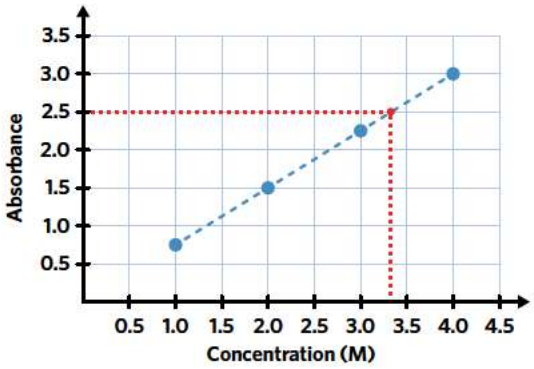


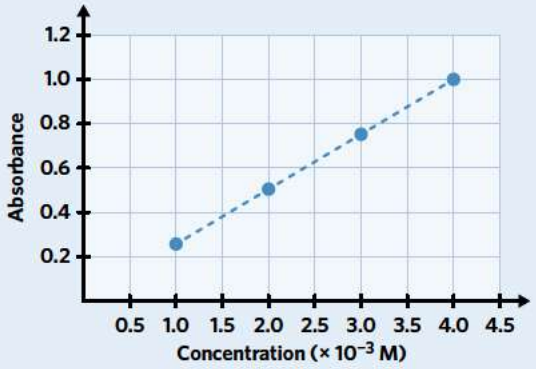
Figure 8 Calibration curve.

The following steps are used to accurately perform colorimetric analysis of an unknown substance:

1. Prepare a set of standard solutions of the substance at known concentrations.
2. Measure the absorbance of the standard solutions at a particular wavelength of visible light.
3. Plot a calibration curve using the collected data, with concentration on the x-axis and absorbance on the y-axis.
4. Measure the absorbance of the solution of unknown concentration using a colorimeter.
5. Determine the concentration of the solution using the calibration curve.

Worked example 1

A researcher is using colorimetry to determine the amount of copper ions in a water sample. Having already constructed a calibration curve, the absorbance of orange light passing through the sample is measured, with a result of 0.62. Using the calibration curve shown, determine the concentration of copper ions in the water sample.



What information is presented in the question?

The absorbance of the sample and the calibration curve.

What is the question asking us to do?

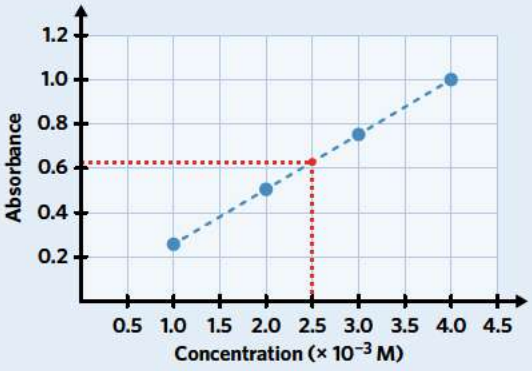
Determine the concentration of copper ions in the water sample.

What strategy(ies) do we need in order to answer the question?

1. Use the absorbance value to plot a point on the calibration curve.
2. Determine the concentration of the sample by observing the x-axis value.

Answer

The absorbance of the sample is 0.62.



Based on the calibration curve, it can be seen that the concentration of copper ions is 2.5×10^{-3} M.

When using a calibration curve to determine the concentration of a solution, the absorbance value obtained must not be greater than that of the standard solution with the highest concentration. If the absorbance value obtained is greater than the maximum value or lower than the minimum value on the calibration curve, the curve cannot be used to accurately determine the concentration of the solution, therefore more standards need to be measured.

Analysis of water using UV-visible spectroscopy 2.2.10.2

OVERVIEW

UV-visible spectroscopy is an analytical technique using ultraviolet and visible light to determine the quality of a sample of water.

THEORY DETAILS

Another form of chemical analysis using many of the same principles as colorimetry is **UV-visible spectroscopy**. This technique is more accurate than others we have explored due to the use of a more sophisticated instrument – a **UV-visible spectrophotometer**. This instrument uses a **monochromator** to select the specific wavelength of light to be used in spectroscopic analysis.

The solution being analysed is first scanned across multiple wavelengths by the spectrophotometer to assess which wavelengths are absorbed to the greatest extent by the solution. The wavelength that is most absorbed by the solution is the one that is selected for spectroscopy.

In the scan shown in figure 9, strong absorbance is shown at a wavelength of 250 nm, which is in the ultraviolet range of the electromagnetic spectrum, as well as a smaller peak around 525 nm, which corresponds to the colour green in the visible light region. Generally, the analysis would be conducted at the wavelength corresponding to the highest absorbance however we also need to make sure that it does not also have significant absorbance from other substances dissolved in the solution. Therefore, the wavelength chosen may not be the one with the strongest absorbance by the solution. For the scan shown in figure 9, the wavelength of 525 nm may be used if other substances in the solution also readily absorb light with a wavelength of 250 nm.

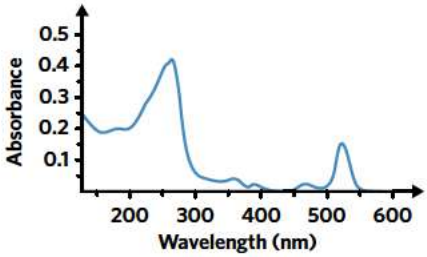
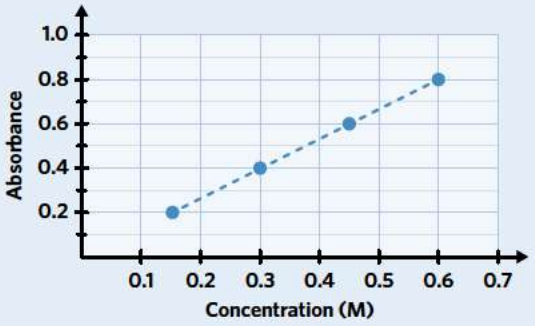


Figure 9 Spectrophotometer scan.

Like colorimetry, UV-visible spectroscopy involves the construction of a calibration curve from the absorbance of a series of standard solutions. The absorbance from the solution of unknown concentration can be compared to the curve to determine the concentration of the solution.

Worked example 2

UV-visible spectroscopy makes use of calibration curves to visualise the relationship between concentration and absorbance of light in solutions. A student produces the following calibration curve, and determines the absorbance of a solution with unknown concentration to be 0.7. Determine the concentration of the solution.



What information is presented in the question?

The absorbance of the sample and the calibration curve.

What is the question asking us to do?

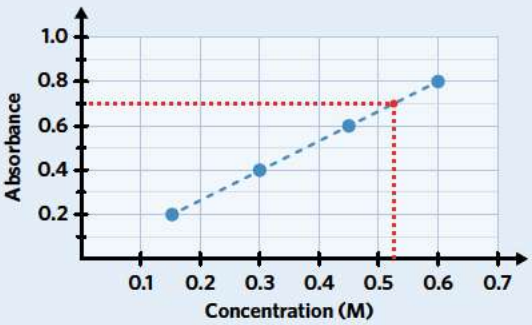
Determine the concentration of the solution.

What strategy(ies) do we need in order to answer the question?

1. Use the absorbance value to plot a point on the calibration curve.
2. Determine the concentration of the sample by observing the x-axis value.

Answer

The absorbance of the sample is 0.7.



Based on the calibration curve, it can be seen that the concentration of the solution is 0.525 M.

Metal complexes

Some metal ions do not absorb UV or visible light effectively. Hence, they require the formation of a metal complex in order to be observed in spectroscopy. For example, solutions containing $Fe^{3+}_{(aq)}$ ions neither absorb visible nor ultraviolet light effectively. However, in a complex such as $FeSCN^{2+}_{(aq)}$, the solution is brightly coloured and thus suitable for UV-visible spectroscopic analysis. This process is shown in figure 10.

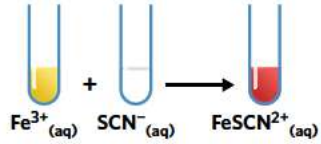


Figure 10 $FeSCN^{2+}$ metal complex.

Uses of colorimetry and UV-visible spectroscopy

Both colorimetry and UV-visible spectroscopy are extremely useful in determining the concentration of solutes in a given solution. This has a number of real-world applications, including but not limited to:

- Identifying and quantifying the contents of blood including bilirubin, cholesterol, iron, haemoglobin and sugars.
- Assessing the contents of urine such as lead.
- Carrying out quality assurance of milk and other dairy products.
- Testing air for the presence of chromium and lead.
- Assessing the phosphate concentration in waterways to prevent eutrophication.

Theory summary

- Colorimetry makes use of the interactions between light and ions to determine water sample quality.
- The complementary colour of that most strongly absorbed will be the colour most strongly reflected.
- UV-visible spectroscopy uses many of the same principles as colorimetry but with more sophisticated equipment.

10D QUESTIONS

Theory review questions

Question 1

The electromagnetic spectrum includes (but is not limited to)

A x-rays, visible light and ultraviolet radiation.

B gamma rays, radio waves and sound waves.

Question 2

The complementary colour of the wavelength of light most strongly absorbed is

A strongly absorbed.

B reflected.

Question 3

A calibration curve is used in colorimetry to

A determine the concentration of an ion with known absorbance of light.

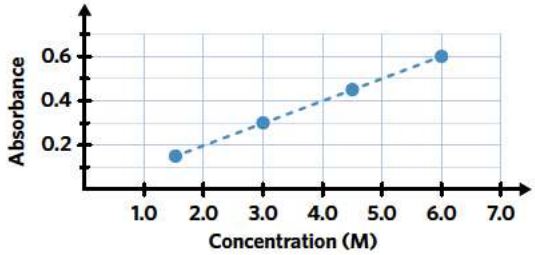
B visualise the inverse relationship between ion concentration and light absorbance.

Question 4

What is the concentration of a substance with an absorbance of 0.25?

A 1.25 M

B 2.50 M



Question 5

The colour of an object corresponds to

A the colour most strongly absorbed by the material.

B the colour reflected or transmitted through the material.

Question 6

The benefit of a spectrophotometer over regular colorimeters is that it is

A able to produce an emission spectrum.

B able to scan different light wavelengths to assess which has the greatest absorbance.

Question 7

A monochromator is used in UV-visible spectroscopy to

A reduce the intensity of light being used.

B select a particular wavelength of light.

Question 8

Calibration curves are constructed through

A measuring the absorbance of solutions of known concentration.

B measuring the concentrations of solutions with known absorbance values.

Question 9

One potential use of colorimetry and UV-visible spectroscopy is

A measuring electrical conductivity of a water sample.

B assessment of blood contents.

Question 10

Metal complexes can be used for

A solutions that do not absorb light to enable absorption.

B solutions that readily absorb light to reduce the intensity of absorption.

Deconstructed questions

Use the following information to answer questions 11–13.

A chemistry teacher is teaching a class about spectroscopy, most notably colorimetry and UV-visible spectroscopy, and has been asked a series of questions.

Question 11 (1 MARK)

Which option correctly describes colorimetry?

A Measuring the absorbance of light passing through a solution to establish the concentration of ions.

B Inserting a sample into a flame and observing the colour produced to determine the chemical composition of the sample.

C Determining the amount of a substance dissolved in water by weighing the precipitate formed.

D Producing an emission spectrum to determine the presence of particular ions in a solution.

Question 12 (1 MARK)

Which option correctly describes the difference between colorimetry and UV-visible spectroscopy?

A Ultraviolet electromagnetic radiation is used in UV-visible spectroscopy instead of visible light.

B More advanced equipment is used in UV-visible spectroscopy for more accurate results.

C More advanced equipment is used in colorimetry for more accurate results.

D UV-visible spectroscopy produces an emission spectrum while colorimetry does not.

Question 13 (3 MARKS)

Explain how colorimetry and UV-visible spectroscopy make use of the interactions between ions in solution and light to provide information about a sample.

Exam-style questions

Within lesson

Question 14 (10 MARKS)

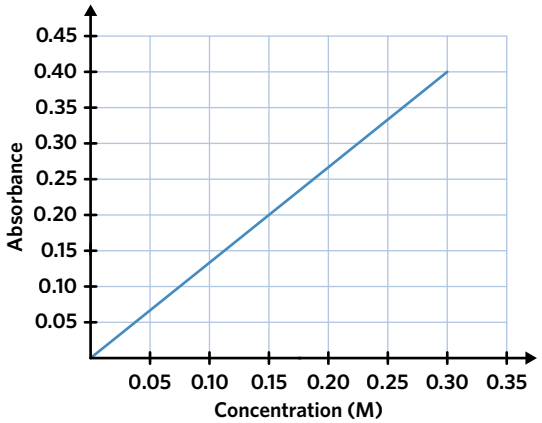
Colorimetry is a form of chemical analysis used to determine the elements present in a solution through their interactions with electromagnetic radiation. In order to determine the concentration of these elements, a calibration curve must be consulted.

a What is a calibration curve? (1 MARK)

b Describe the process used to produce a calibration curve. (3 MARKS)

c Describe how a calibration curve is used to determine the concentration of a sample solution. (2 MARKS)

d The calibration curve shown was constructed using standard solutions of copper ions in water. Using the calibration curve, determine the approximate concentration of a solution of copper ions with an absorbance of 0.25. (1 MARK)



e A solution of copper ions has a blue colour. What colour light will be used for colorimetry? Explain your reasoning. (3 MARKS)

Question 15 (10 MARKS)

Ryan has recently learned about colorimetry in his chemistry class, and wants to put his knowledge to use. At the school laboratory, Ryan produces a series of standard solutions of cobalt chloride and measures the absorbance of each.

Concentration	Absorbance
0.5 M	0.19
1.0 M	0.40
1.5 M	0.60
2.0 M	0.81

a What is the unit used for absorbance in colorimetry? (1 MARK)

b Produce a calibration curve using Ryan's experimental results. Include correctly labelled axes and units. (3 MARKS)

c Ryan claims that violet light should be used to carry out colorimetry as the cobalt chloride solution is violet. Explain whether Ryan is correct. (3 MARKS)

d Ryan wishes to use his calibration curve to determine the concentration of another solution containing cobalt chloride.

i The absorbance of the unknown solution is 0.7. Using the calibration curve, determine the concentration of the solution. (1 MARK)

ii Ryan uses colorimetry to determine the absorbance of yet another unknown solution, which has an absorbance of 0.90. Explain why this calibration curve cannot be used for this solution. (2 MARKS)

Question 16 (8 MARKS)

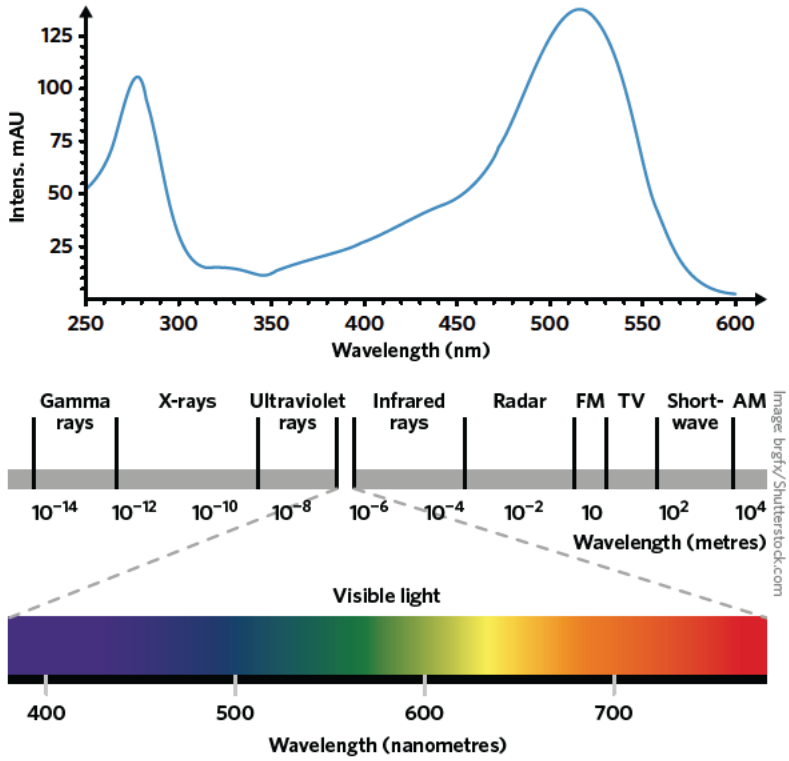
UV-visible spectroscopy is a form of chemical analysis that uses many of the same principles as colorimetry, observing the interactions between light and molecules in solution to determine concentration.

a A UV-visible spectrophotometer uses a monochromator for spectroscopy. What is the function of a monochromator, and what is used instead in colorimetry? (2 MARKS)

b A UV-visible spectrophotometer has the ability to scan a number of different wavelengths of light in order to assess for absorbance.

i Explain why a UV-visible spectrophotometer is advantageous for UV-visible spectroscopy. (2 MARKS)

ii Using the following UV-visible spectrum and electromagnetic spectrum, determine what colour would be best suited for this solution. (1 MARK)



c Spectroscopy relies on the interactions between ions in a solution and light, which requires the solution to be coloured. Describe the process utilised to perform UV-visible spectroscopy of a sample of metal ions without colour. (2 MARKS)

d Identify a real-world application of UV-visible spectroscopy. (1 MARK)

Question 17 (9 MARKS)

Mark has a solution of iron ions with an unknown concentration and is debating between colorimetry and UV-visible spectroscopy.

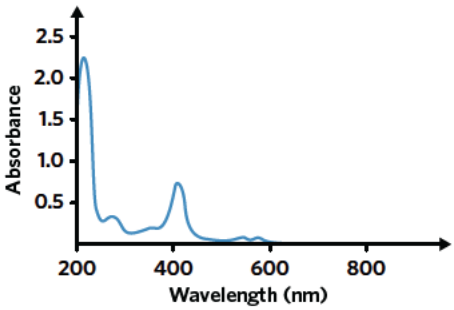
a Explain one advantage and one disadvantage of colorimetry and UV-visible spectroscopy. (4 MARKS)

b Mark decides to use UV-visible spectroscopy for chemical analysis, but wants to clarify his understanding of the method first. Using the UV-visible spectrum, identify which of the following statements are true, justifying your answer(s). (3 MARKS)

I The spectrum is produced by measuring the frequencies of light from electrons returning to the ground state from an excited state.

II The concentration of a solution of this compound can only be determined by UV-visible spectroscopy at 210 nm.

III The amount of light absorbed by a solution depends on its concentration.



c The iron solution is transparent. Explain how light used in the colorimetric analysis will interact with the solution. (2 MARKS)

Multiple lessons

Question 18 (7 MARKS)

Annie mixed two test tubes, each containing silver and iodide ions at different concentrations. She then uses the final solution in a colorimetry test to determine the concentration of silver ions in the solution.

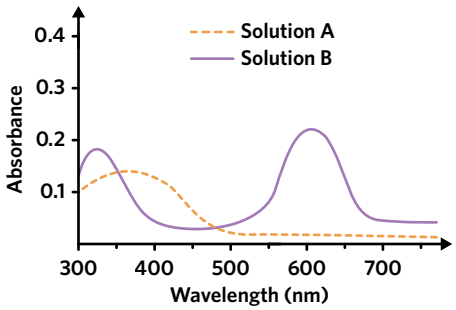
a Comment on the appropriateness of using colorimetry to measure the concentration of silver ions. (3 MARKS)

b As part of the experiment, Annie mixed 20.0 mL of a 1.0 M silver ion solution with 15.0 mL of a 1.2 M iodide ion solution. Determine the limiting reagent. (3 MARKS)

c What amount, in mol, of silver ions would be available in the solution to absorb light as part of the experiment? (1 MARK)

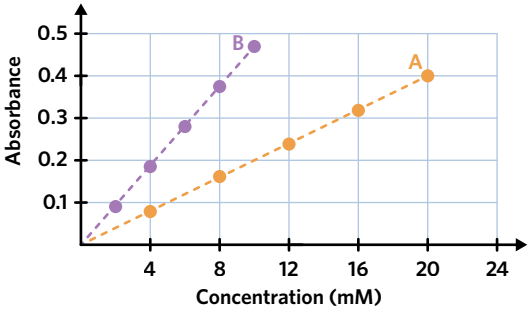
Question 19 (7 MARKS)

UV-visible spectroscopy was used to measure the spectra of solutions A and B. Solution A was orange, and solution B was violet. The following UV-visible spectra were recorded.



a If 10.00 mL of solution A was mixed with 10.00 mL of solution B, which wavelength should be used to measure the absorbance of solution B? Justify your answer. (2 MARKS)

b The analyst used two sets of standard solutions and blanks to determine the calibration curves for the two solutions. The absorbances were plotted on the same axes, as shown on the graph.



The analyst found that when measured at an appropriate wavelength, solution B had an absorbance of 0.21. If solution B was cobalt(II) nitrate, $Co(NO_3)_2$, determine it's concentration (M). (2 MARKS)

c A small sample of cobalt(II) nitrate was adde to copper sulfate.

i Write the balanced equation for this reaction. (1 MARK)

ii What concentration of a 20.0 mL sample of copper sulfate needs to be mixed with 10.0 mL of cobalt(II) nitrate so that the same amount of both substances is added to the reaction? (2 MARKS)

Key science skills

Question 20 (5 MARKS)

A student is analysing a sample of water that contains copper ions using colorimetry. The student sets up a series of standard solutions in a number of test tubes, some of which are transparent and some of which are translucent, and then carries out colorimetric analysis. When she constructs the calibration curve after the experiment, she discovers that it does not show the expected linear shape.

a Explain why the calibration curve does not fit the expected linear shape. (2 MARKS)

b What type of error is this? Explain. (2 MARKS)

c What should the student do to remove the source of error? (1 MARK)

Exam-style question hints

14a: Calibration curves are produced from standard solutions. **14b:** Calibration curves are produced from standard solutions. **14c:** Calibration curves show concentration on the x-axis, and absorbance on the y-axis. **14d:** Every point on the curve has an x-value and a y-value. **14e:** Take ROYGBV, and shift it 3 letters to the right. **15a:** Some properties do not have units. **15b:** Calibration curves have concentration on the x-axis, and absorbance on the y-axis. **15c:** Colorimetry makes use of complementary colours. **15d:** Every point on the curve has an x-value and a y-value. **16a:** A monochromator can increase the accuracy of spectroscopy. **16bi:** Spectrophotometers make absorbance take place more effectively. **16bii:** Strong absorbance is important for carrying out spectroscopy. **16c:** Ionic solutions can become coloured. **16d:** UV-visible spectroscopy can be used to analyse the contents of substances. **17a:** Consider factors such as cost and accuracy. **17b:** A UV-visible spectrum shows which wavelengths of light have the highest absorbance. **17c:** Different wavelengths of light have different interactions with ions. **18a:** Gravimetric analysis involves the addition of a reactant. **18b:** The limiting reagent is completely used up in the reaction. **18c:** This value is equal to the number of mol of silver ions left over at the end of the reaction. **19a:** The wavelength used to measure the absorbance of a solution is the wavelength at which the solution absorbs the most light without interference from other solutions or contaminants. **19b:** Every point on the curve has an x-value and a y-value. **19c:** Consider how cobalt (II) nitrate reacts with copper sulfate. **20a:** Calibration curves should depict a linear relationship if completed properly. **20b:** Errors vary depending on if they affect results consistently or not. **20c:** Variables must be controlled in experiments.

10E ANALYSING WATER USING LIGHT - PART 2

In this lesson, we will be learning about two key methods of chemical analysis – atomic emission spectroscopy and atomic absorption spectroscopy.

10A Water on Earth | 10B Water contamination | 10C Analysing water using gravimetric analysis | 10D Analysing water using light - part 1 | 10E Analysing water using light - part 2 | 10F Volumetric analysis | 10G Chromatography

Study design dot point

- the application of atomic absorption spectroscopy (AAS), including the use of a calibration curve, to determine the concentration of metals or metal ions in a water sample (excluding details of instrument)

Key knowledge units

Principles of AES and AAS	2.2.11.1
Determining the concentration of metal ions using a calibration curve	2.2.11.2

Key terms and definitions

Excited electron electron which has absorbed energy and is elevated to a higher unstable energy level

Ground state lowest energy level that an electron occupies

Atomic emission spectroscopy (AES) method of chemical analysis using the intensity of light emitted from a flame to determine the metals present

Emission spectrum spectrum of electromagnetic wavelengths emitted by a metal ion in atomic emission spectroscopy

Atomic absorption spectroscopy (AAS) method of chemical analysis using the light absorbed by metals to determine the type and number of metals present in a sample

Lesson links

This lesson builds on:

- 2A The atom

Interactions between electrons and energy are observed in spectroscopy.

Principles of AES and AAS 2.2.11.1

OVERVIEW

Flame tests, atomic emission spectroscopy and atomic absorption spectroscopy are chemical analysis techniques that are commonly used to determine the quality and quantity of metals in a sample.

THEORY DETAILS

We have already covered a range of different methods that can be used to carry out chemical analysis. In this lesson, we will be exploring another pair of chemical analysis techniques, namely atomic emission spectroscopy and atomic absorption spectroscopy. These analytical techniques involve observing changes in the energy levels of electrons to obtain qualitative and quantitative information about a sample.

Flame tests

As we learned in lesson 2A, electrons are able to jump from a lower energy state to an excited state. This can occur when the substance is exposed to a flame, which energises an electron to become an **excited electron**. However, excitation will only take place if the amount of energy absorbed is precisely equal to the difference in energy between the lower and higher energy levels.

Electrons are unstable in the energised state. Consequently, electrons rapidly return to their original, lower energy state termed the **ground state**. When this occurs, the energy that was absorbed by electrons to elevate them to the higher energy state will be emitted as electromagnetic radiation.

When metals are heated, they emit colour due to excitation of electrons in the metal sample. For example, the exposure of table salt (NaCl) to an open flame will colour the flame yellow. Different elements require different amounts of energy to excite electrons. The consequence of this is that each metal will emit light of a different wavelength when returning to the ground state which appears as a unique colour. This is shown in figure 1. The specific wavelength emitted can be recorded and used to identify metals making up a substance.

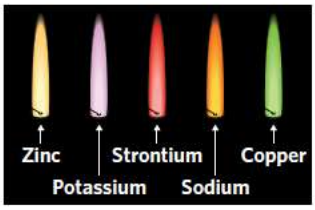


Image: Zern Liew/Shutterstock.com

Figure 1 Flame test colours.

The flame test can be used to identify the metal ions in a substance by introducing a sample of the substance to a non-luminous (non coloured) bunsen burner flame as shown in figure 2. This is the flame produced when the air-hole is open. By conducting a flame test, the colour produced by the metal ion in an unknown sample can be compared to the characteristic colour produced by known samples of metals. Since each flame colour is unique, the colour can be used to identify particular metals in a sample.

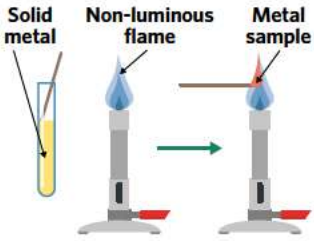


Figure 2 Flame test.

There are several limitations of the flame test method. A significant issue is that flame tests cannot be used for metal ion identification if the particular species does not produce a coloured flame. Furthermore, while each metal ion's colour is unique, it may be difficult to distinguish between, for example, different shades of blue. There may also be contaminants in the sample masking the colour emitted by the ions we wish to analyse.

Atomic emission spectroscopy

While flame tests provide basic information about the ionic composition of a sample, as mentioned previously, there are several deficiencies associated with the method. To avoid the issues associated with flame tests, an alternative technique is employed known as **atomic emission spectroscopy (AES)**. An atomic emission spectrometer is the instrument used for atomic emission spectroscopy. This apparatus operates in a similar way to that used for flame tests. However, unlike in flame tests, the spectrometer enables the identification of all ions present in a sample.

There are two fundamental differences between the flame test and atomic emission spectroscopy which makes the latter a more accurate technique:

1. The flame used in atomic emission spectroscopy operates at a higher temperature compared to the bunsen burner flame utilised in flame tests. Consequently, there is sufficient energy to excite electrons in a greater number of elements.
2. In atomic emission spectroscopy, light is passed through a prism to separate the light into its constituent wavelengths. This produces a display of lines with each line representing a wavelength of light. This is known as an **emission spectrum.**

The process involved in AES is shown in figure 3.

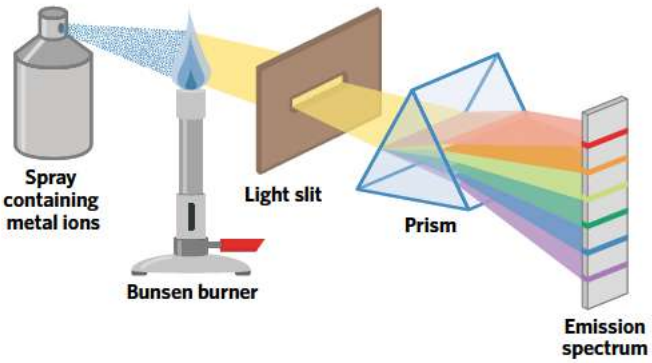


Figure 3 Atomic Emission Spectroscopy.

Since the chemical makeup of every element is different, the energy transitions in the process of electron excitement vary for different atoms. Consequently, each element will have a unique emission spectrum, as demonstrated in figure 4 with hydrogen, helium and mercury each having distinct emission spectra. These spectra operate as a 'fingerprint' and hence can be used to identify the different elements present in a substance. Further investigation into the intensity of the light at the different wavelengths can also give quantitative information about the element(s) in the sample, however emission spectrums themselves only give us qualitative data.

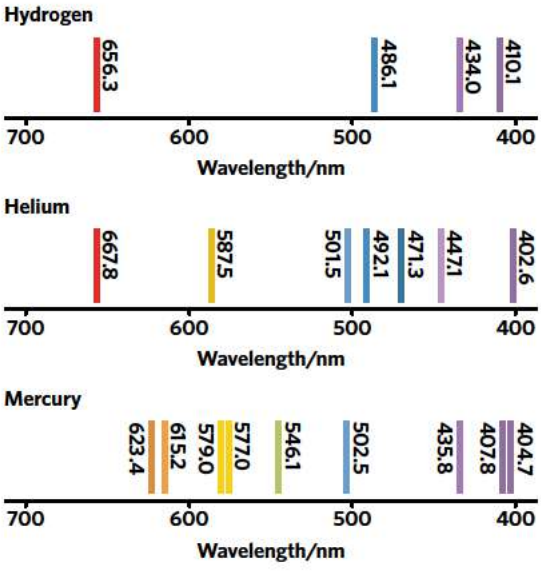


Figure 4 Emission spectra.

Atomic absorption spectroscopy

Another form of chemical analysis, similar to atomic emission spectroscopy, is **atomic absorption spectroscopy (AAS)**. This technique is widely used to determine the presence and concentration of metals in substances. Hence, AAS is both a qualitative and quantitative technique.

AAS works on the principle that, much like the heat energy in the flame test, metal atoms absorb light and their electrons become excited if the energy of the light is exactly equal to the energy difference between an electron's ground state and excited state. Hence, the amount of light absorbed at a specific wavelength can be used to determine the amount of a specific element in a sample. This is because each element has a unique pattern of absorption. As the name suggests, atomic absorption spectroscopy differs from atomic emission spectroscopy on the basis that the former uses absorbance rather than emission data.

The method of atomic absorption spectroscopy is detailed as follows:

1. A hollow cathode lamp, composed of the metal under analysis, emits specific wavelengths of light.
2. Light produced by the lamp passes through, and is absorbed by, a sample that is vaporised in a flame.
3. A monochromator selects a particular wavelength of light. The wavelength chosen for analysis depends on the element being tested.
4. A detector determines how much of the particular wavelength of light is passed through the sample. The absorbance of the wavelength of light is used to measure the concentration of a particular element in the sample.

This process is shown in figure 5. It is important to note that this process must be repeated for each metal present since the technique is only able to identify and quantify elements individually.

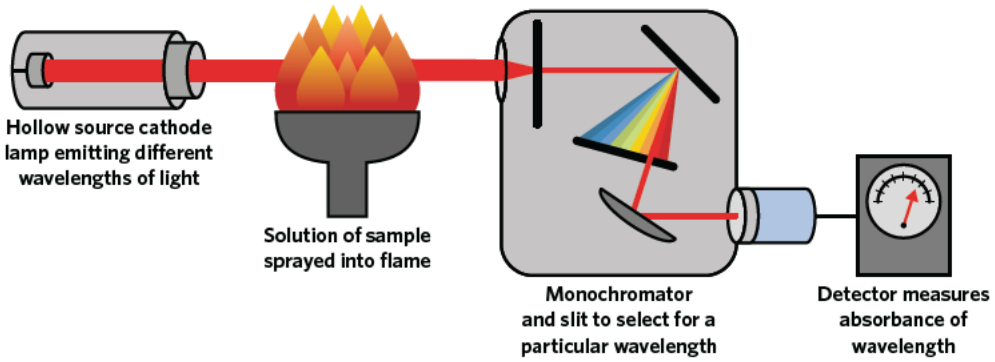


Figure 5 Atomic absorption spectroscopy.

For example, if a sample was to be analysed for sodium ion concentration, then a sodium lamp would be used for AAS. The sample is vapourised in the flame, and the frequencies of light produced by the sodium lamp would pass through the vapourised sample in the flame. Only the sodium ions in the sample would absorb those wavelengths and this would be recorded by the detector.

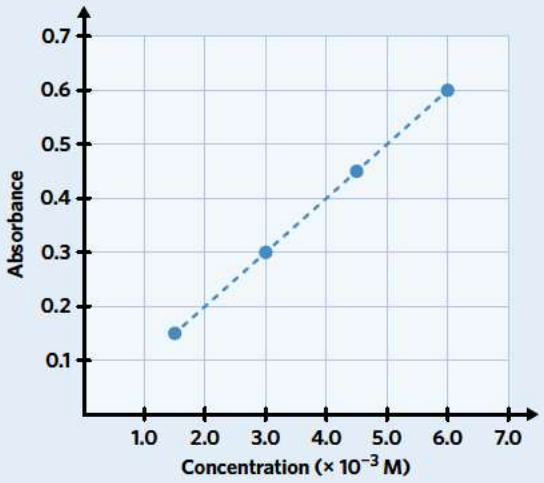
Determining the concentration of metal ions using a calibration curve 2.2.11.2

AAS calibration curves

To find the concentration of a particular metal ion in a sample, a calibration curve must be created. Much like the calibration curves introduced in colorimetry earlier in the chapter, constructing an AAS calibration curve requires a series of standard solutions of precisely known concentrations containing the metal ion under analysis. By performing AAS on the standard solutions and plotting the absorbance data on a graph, the concentration of the metal ion in the sample can be established using the obtained absorbance value and the calibration curve. This information can be obtained since the relationship between concentration and absorbance is linear.

Worked example 1

A solution with an unknown concentration of copper was found to have an absorbance value of 0.4. Using the copper calibration curve below, measure the concentration of copper in the solution.



What information is presented in the question?

The absorbance of copper and the copper calibration curve.

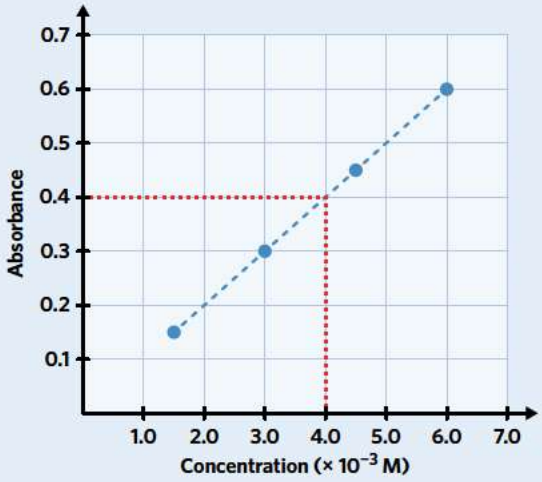
What is the question asking us to do?

Measure the concentration of copper in the solution using the calibration curve.

What strategy(ies) do we need in order to answer the question?

1. Find the point on the curve depicting an absorbance of 0.4.
2. Find the x-value of the curve at the point where absorbance is equal to 0.4 to determine the concentration of copper in the solution.

Answer



The concentration of copper in the solution at the point where the absorbance is equal to 0.4 is 4.0×10^{-3} M.

To ensure that the concentrations derived from calibration graphs are accurate, we need to make sure that the absorbance value of the sample falls well within the range of the standards used to create the calibration graph.

Theory summary

- The flame test identifies elements in a sample by observing light emission upon electrons returning to the ground state following excitation.
- Atomic emission spectroscopy (AES) uses the same principles as the flame test but produces an emission spectrum.
- Atomic absorption spectroscopy (AAS) measures how much of a particular wavelength of light has been absorbed by a sample to determine its concentration.

10E QUESTIONS

Theory review questions

Question 1

The flame test produces flames with different colours due to

A differences in the amount of energy required to excite electrons to a higher energy level.

B different wavelengths of light being absorbed by different materials in the flame.

Question 2

The ground state of an electron

A is considered unstable as the electron is easily excitable.

B is the lowest possible energy state of a given electron.

Question 3

An excited electron is

A an electron that has been energised and elevated to a higher energy level.

B an electron that has recently released energy.

Question 4

The primary difference between atomic emission spectroscopy and atomic absorption spectroscopy is that

A AAS measures the wavelength of light absorbed whereas AES measures the wavelength of light emitted.

B AAS produces an emission spectrum whereas AES doesn't.

Question 5

Atomic emission spectroscopy is more effective than the flame test because

A AES doesn't require the use of a flame.

B AES passes light through a prism to produce an emission spectrum.

Question 6

An emission spectrum

A is the splitting of light due to it passing through a prism.

B is used to identify excited and ground state electrons.

Question 7

Calibration curves used in atomic absorption spectroscopy

A have a non-linear shape.

B visualise the relationship between concentration and light absorbance.

Question 8

A monochromator is used in atomic absorption spectroscopy to

A select a particular wavelength of light for analysis.

B reduce the intensity of light for the detector.

Question 9

An emission spectrum produced in atomic emission spectroscopy

A is unique for each element.

B can be inaccurate due to similarities between elements.

Question 10

Atomic absorption spectroscopy involves

A heating a sample and observing the emission spectrum produced.

B passing a specific wavelength of light through a sample to determine the concentration of particular elements.

Deconstructed questions

Use the following information to answer questions 11–13.

Sufjan is a chemist attempting to analyse a sample of a solution containing titanium ions. He wishes to use a form of spectroscopy, and is debating between the use of atomic emission spectroscopy and atomic absorption spectroscopy.

Question 11 (1 MARK)

Choose the option that best describes the method used in atomic emission spectroscopy.

A Assessing the wavelength of light most strongly reflected by a sample.

B Observing how much of a wavelength of light is absorbed by a sample.

C Observing the emission spectrum produced from electrons transitioning from an excited state to ground state.

D Weighing the precipitate formed between titanium and a reactant.

Question 12 (1 MARK)

Choose the option that best describes the method used in atomic absorption spectroscopy.

A Assessing the wavelength of light most strongly reflected by a sample.

B Observing how much of a wavelength of light is absorbed by a sample.

C Observing the emission spectrum produced from electrons transitioning from an excited state to ground state.

D Weighing the precipitate formed between titanium and a reactant.

Question 13 (3 MARKS)

Sufjan claims that the flame test could be a suitable replacement for atomic emission spectroscopy. Comment on the accuracy of this statement, and justify your answer.

Exam-style questions

Within lesson

Question 14 (1 MARK)

Which of the following is accurate regarding atomic absorption spectroscopy and atomic emission spectroscopy?

A AAS cannot produce an emission spectrum as insufficient energy is being put into the sample.

B AES uses a calibration curve to visualise the relationship between concentration and absorbance.

C AAS observations are based on the return of electrons from an excited state to the ground state.

D AES can be used for qualitative analysis.

Question 15 (1 MARK)

Tim is a year 11 chemistry student interested in carrying out chemical analysis on a magnesium sample using atomic absorption spectroscopy. Which of the following setups would be most appropriate?

A Passing light from a hollow sodium cathode lamp through a solution in a test tube and measuring the absorbance of a particular wavelength of light with a detector.

B Passing light from a hollow magnesium cathode lamp through a flame with the sample sprayed in it and observing the emission spectrum produced from it.

C Passing light from a hollow magnesium cathode lamp through a flame with the sample sprayed in it and measuring the absorbance of a particular wavelength of light with a detector.

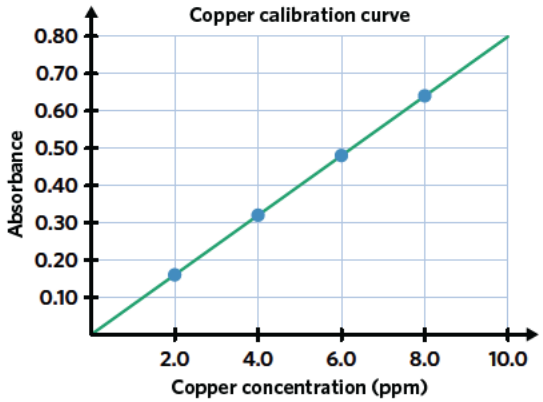
D Heating the sample with a flame and observing the colour produced.

Question 16 (1 MARK)

A student produces the following calibration curve for a solution of copper with an unknown concentration using an atomic absorption spectrometer.

The range that would enable the concentration of the solution to be measured most accurately is

A 2.0 to 8.0 ppm.
B 0 to 10.0 ppm.
C 4.0 to 6.0 ppm.
D 0 to 8.0 ppm.



Question 17 (1 MARK)

Which of the following substances would best be analysed using atomic absorption spectroscopy?

A Iodine
B Nitrogen
C Sodium
D Hydrogen

Question 18 (8 MARKS)

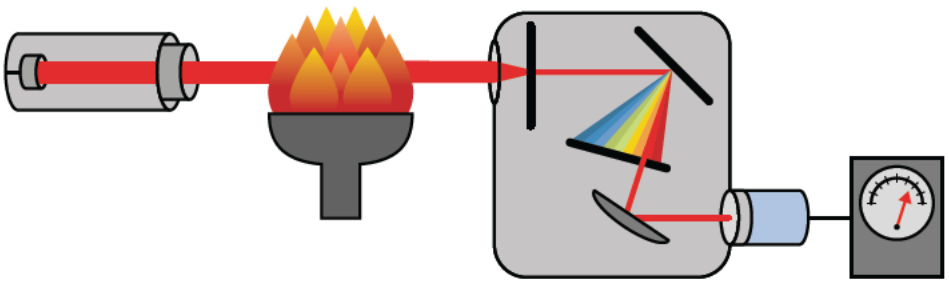
Atomic absorption spectroscopy (AAS) and atomic emission spectroscopy (AES) utilise the interactions of electrons with energy to perform quantitative and/or qualitative spectroscopic analysis. The principles of these forms of analysis are similar to those of the flame test.

a Explain how AES utilises similar principles to the flame test and the advantages of AES. (3 MARKS)

b Determine whether emission spectra can be used for qualitative or quantitative analysis. (1 MARK)

c Explain how the processes differ for AAS and AES and the type of information that is acquired from each technique. (3 MARKS)

d Label the following diagram with the name and function of each instrument in atomic absorption spectroscopy: (1 MARK)



Multiple lessons

Question 19 (9 MARKS)

A chemistry professor plans to construct an AAS calibration curve by obtaining a number of standard solutions containing copper sulphate and measuring the absorbance of light through each.

a Define the term 'solubility'. (1 MARK)

b Explain why copper sulphate is able to dissolve in water. Describe this process. (3 MARKS)

c Construct a calibration curve using the following values. (2 MARKS)

Concentration of copper (mg L^{-1})	1.0	2.0	3.0	4.0	Sample
Absorbance	0.061	0.119	0.178	0.243	0.149

d What is the approximate concentration of the sample? (1 MARK)

e Copper sulphate solution is often used to eradicate parasites from aquariums, but must be at a certain concentration to perform this function. The sample of copper sulphate solution is from the professor's aquarium where he had used copper sulphate for this purpose.

The required concentration for copper sulphate to be effective as an anti-parasitic is 1200 ppb. Is the concentration of the professor's aquarium sufficient? Show your working. (2 MARKS)

Key science skills

Question 20 (6 MARKS)

Carrie is a chemistry student performing atomic absorption spectroscopy. To determine the concentration of copper in an unknown sample, she produces a series of standard solutions and performs atomic absorption spectroscopy on each to record their absorbance and produce a calibration curve. She records the following data:

Concentration	0.25 M	0.50 M	0.75 M	1.00 M
Absorbance	0.14	0.27	0.30	0.56

a Which of these results is unexpected? Explain your answer. (2 MARKS)

b Suggest at least one reason why the error identified in q20a may have taken place. (1 MARK)

c Provide at least three variables that must be controlled when performing atomic absorption spectroscopy. (3 MARKS)

Exam-style question hints

Q14: Consider the difference between emission and absorption. **Q15:** AAS measures the light absorbed by a substance. **Q16:** Values outside of a determined range require extrapolation. **Q17:** AAS is the preferred method for a particular type of element. **Q18a:** The colour of a flame provides information for identification. **Q18b:** AES produces an emission spectrum. **Q18c:** AAS analyses the wavelength of light most strongly absorbed. **Q19a:** Solubility refers to dissolution characteristics. **Q19b:** Dissolution occurs due to interactions between the solute and solvent. **Q19c:** Calibration curves have concentration on the x-axis, and absorbance on the y-axis. **Q19d:** The calibration curve is developed to determine the concentration of unknown substances. **Q19e:** Any concentration less than this will not be effective. **Q20a:** Calibration curves should have a linear shape. **Q20b:** Changes to the method may alter the results. **Q20c:** Only the factor being assessed, the independent variable, should be manipulated.

10F VOLUMETRIC ANALYSIS

In this lesson, we will learn an analytical technique called titration and its role in determining the concentration of an unknown solution.

10A Water on Earth	10B Water contamination	10C Analysing water using gravimetric analysis	10D Analysing water using light - part 1	10E Analysing water using light - part 2	10F Volumetric analysis	10G Chromatography

Study design dot point

- volume-volume stoichiometry (solutions only) and application of volumetric analysis including the use of indicators, calculations related to preparation of standard solutions, dilution of solutions and use of acid-base titrations to determine the concentration of an acid or a base in a water sample

Key knowledge units

Volume-volume stoichiometry	2.2.15.1
Principles of titrations	2.2.15.2
Acid-base titrations	2.2.15.3

Key terms and definitions

Concentration amount of chemical substance in a unit of volume

Volume amount of space occupied by a substance

Dilution process of decreasing concentration of a solution by adding more solvent, usually water

Dilution factor ratio of volume of stock solution to the volume of final (diluted) solution

Volumetric analysis quantitative analytical technique used to determine unknown concentrations of solutions

Titration quantitative technique used to find the unknown concentration and/or amount of substance in a solution

Titrant solution of known concentration used in a titration reaction

Primary standard solution solution whose concentration can be precisely calculated

Primary standard substance used to make a primary standard solution

Pipette laboratory tool to measure and deliver small volumes of solutions accurately

Aliquot an exact amount of a sample solution that is to be analysed during titration

Conical flask glassware which the analyte is placed in

Burette glassware which delivers accurate, yet variable volumes of titrant

Meniscus curved upper surface of liquid

Titre volume of the solution delivered from the burette to reach the end point of a titration

Concordant titres volume of three or more titres that fall within approximately 0.1 mL of each other

Average titre mean value of the concordant titres

Indicator substance that will experience a colour change under a specific condition

End point point at which a colour change occurs during titration

Equivalence point point at which the reaction has occurred as per the stoichiometric ratio of the balanced equation

Lesson links

This lesson builds on:

- 8A An introduction to acids and bases

Some titration reactions involve acid-base reactions.

Volume-volume stoichiometry 2.2.15.1

OVERVIEW

Volume-volume stoichiometry relates the concentrations and volumes of solutions to the amount of chemical present.

THEORY DETAILS

As we learned in lesson 2F, the mole is a fundamental concept in chemistry which defines a certain number of particles. The number of particles present in 1 mol of a substance is equal to Avogadro's constant which is 6.02×10^{23}. Whenever quantitative chemistry is performed, it is important to work with the mole concept as chemicals react in fixed whole number mole ratios according to chemical equations. For example, consider the following acid-base reaction:

$1HCl_{(aq)} + 1NaOH_{(aq)} \rightarrow 1H_2O_{(l)} + 1NaCl_{(aq)}$

1 mole of hydrochloric acid, $HCl_{(aq)}$, reacts with 1 mole of sodium hydroxide, $NaOH_{(aq)}$, to produce 1 mole of water, $H_2O_{(l)}$, and 1 mole of sodium chloride, $NaCl_{(aq)}$.

However, $HCl_{(aq)}$ and $NaOH_{(aq)}$ are not present as pure substances but rather as solutions. In fact, we will almost deal entirely with solutions in VCE chemistry when working with these compounds. This is because they are both able to be dissolved in water as shown in figure 1.

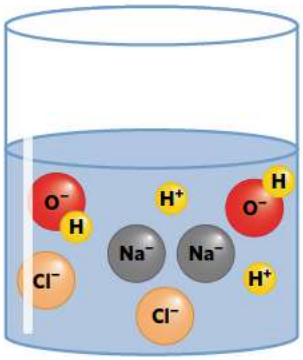


Figure 1 Solute particles dissolved in water.

The number of moles of solute present in solution depends on two factors: concentration and volume. As we learned previously, **concentration** (c) is the amount of chemical substance in a defined **volume** (V). Both of these variables are related the number of mol of particles present by the below equation:

$n = c \times V$

Where n is the number of moles of a substance present, c is the concentration of the solution in mol/L (M) and V is the volume of solution present in litres (L). This equation is found in the VCE Chemistry data book for Units 3 & 4. An important idea to remember is that when calculating the concentration of ions within a compound, we also need to consider the ratios that exist between the compound and the ion itself. For example, when a solution of $Mg(OH)_2$ is fully dissolved in water, there will be one $Mg^{2+}_{(aq)}$ for every two $OH^-_{(aq)}$.

Worked example 1

What is the number of moles of $Cl^-_{(aq)}$ present in 200 mL of 0.25 M solution of $CaCl_{2(aq)}$ solution.

What information is presented in the question?

Volume of solution, V : 200 mL.

Concentration of solution, c : 0.25 M.

What is the question asking us to do?

Find the number of moles of $Cl^-_{(aq)}$ dissolved in the solution.

What strategy(ies) do we need in order to answer the question?

1. Substitute the volume and concentration of $CaCl_{2(aq)}$, using the appropriate units, into the equation $n = c \times V$.
2. Use the dissolution equation $CaCl_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2Cl^-_{(aq)}$ to determine the number of moles of $Cl^-_{(aq)}$ dissolved in the solution.

Answer

$n = c \times V$

$c = 0.25$ M, $V = 0.200$ L

$n(CaCl_2) = 0.25 \times 0.2 = 0.05$ mol

$n(CaCl_2) : n(Cl^-)$

$= 1 : 2$

$n(Cl^-) = \frac{2}{1} \times n(CaCl_2) = 2 \times 0.05 = 0.10$ mol

As we learned previously, concentrations and even volumes can be expressed using different units. Therefore conversions have to take place to the appropriate units before the equation can be used to determine the number of moles of a substance present.

Another complicating factor occurs when a solution is diluted. As we learned in lesson 8B, **dilution** refers to decreasing the concentration of solute particles, usually by adding more solvent. In figure 2, the original solution is diluted by the addition of more solvent.

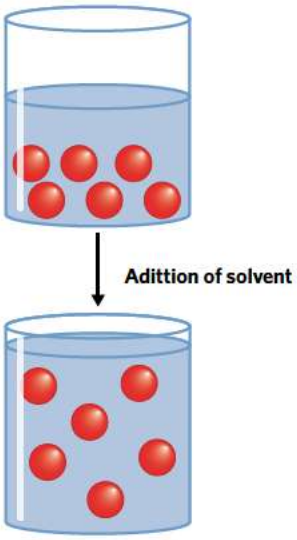


Figure 2 Effect of dilution on a solution.

Since the addition of solvent has not added more particles of solute, the number of moles of solute particles present is the same before and after dilution. However, what has changed is the concentration of particles and volume before and after dilution. Putting this information together:

$n_1 = c_1V_1$

$n_2 = c_2V_2$

$n_1 = n_2$

$c_1V_1 = c_2V_2$

Therefore, given the concentration and volume of the original solution, as well as the change in volume, it is possible to determine the final concentration of the solution after dilution by the above equation.

When working with solutions, the ability to calculate dilutions is an important skill as many chemicals that we use are purchased in large volumes and have very high concentrations. In order for them to be usable in the laboratory, they need to be diluted. For example, consider a laboratory technician who has just purchased a 10 L of 5.0 M HCl solution. However, an experiment requires 1.0 L of 1.0 M HCl solution. The technician therefore has to calculate, using dilution calculations, the extent to which the stock solution would need to be diluted in order to achieve a solution with the appropriate concentration. To do this, the laboratory technician would need to perform the following calculation:

$c_1V_1 = c_2V_2$

$5.0\ \text{M} \times V_1 = 1.0\ \text{M} \times 1.0\ \text{L}$

$$V_1 = \frac{1.0 \times 1.0}{5.0}$$

$= 0.20\ \text{L}$

By comparing the information about the original stock solution against the desired diluted solution, we can see that a total of 200 mL of the 5.0 M stock is required, with the rest of the volume filled with deionised water, to produce a 1.0 L sample of 1.0 M HCl. Alternatively, the technician could determine the amount, in mol, of HCl required in 1.0 L of a 1.0 M instead by the following calculation:

$n(\text{HCl})_{required} = cV = 1.0 \times 1.0 = 1.0\ \text{mol}$

$$V(\text{HCl})_{stock\ solution} = \frac{n}{c} = \frac{1.0}{5.0} = 0.20\ \text{L}$$

As we can see here, both calculations lead to the same result.

Another way of looking at these results is to understand the extent to which the original concentrated solution was diluted. According to our example, we needed to dilute the 200 mL sample to achieve a final volume of 1.0 L. This means that we diluted our sample by a **dilution factor** of :

$$\frac{\text{volume of diluted solution}}{\text{volume of stock solution}} = \frac{1.0\ \text{L}}{0.20\ \text{L}} = 5$$

What this suggests is that the original HCl solution is five times more concentrated than the diluted solution. This will become useful if we ever had to work backwards to find the concentration of an original stock solution using the diluted solution.

This is also important because sometimes solutions have to be diluted before they can be used in a titration. Therefore, to calculate the concentration of the undiluted solution, the concentration of the diluted solution is multiplied by the dilution factor.

Worked example 2

A solution containing 3.0 g of $KOH_{(aq)}$ in 200 mL of water was prepared. It was determined that an experiment required a 500 mL sample of a 0.10 M solution. What volume of the original solution is required to make the 0.10 M 500 mL sample of $KOH_{(aq)}$ solution?

What information is presented in the question?

Mass of $KOH_{(aq)}$: 3.0 g.

Initial volume: 200 mL.

Final volume required: 500 mL.

Final concentration required: 0.10 M.

What is the question asking us to do?

Find the volume of the original $KOH_{(aq)}$ sample required.

What strategy(ies) do we need in order to answer the question?

1. Use the equation $n = \frac{m}{M}$ to find the number of mol of $KOH_{(aq)}$.
2. Calculate the original concentration of $KOH_{(aq)}$ using the equation $c = \frac{n}{V}$.
3. Calculate the volume of the original solution required using the formula $c_1 \times V_1 = c_2 \times V_2$.

Answer

$n = \frac{m}{M}$

$m = 3.0$ g, $M = 39.1 + 16.0 + 1.0 = 56.1$ g mol^{-1}

$n(\text{KOH}) = \frac{3.0}{56.1} = 0.0535$ mol

$c(\text{KOH}) = \frac{0.0535}{\left(\frac{200}{1000}\right)} = 0.267$ M

$c_1 \times V_1 = c_2 \times V_2$

$0.267\ V_1 = 0.10 \times \left(\frac{500}{1000}\right)$

$V_1 = 0.187$ L $= 0.19$ L

Therefore, 0.19 L or 190 mL of the original $KOH_{(aq)}$ solution would be required to make a 500 mL solution of 0.10 M KOH.

Putting this together, we can use the above calculations to generate the flowchart shown in figure 3.

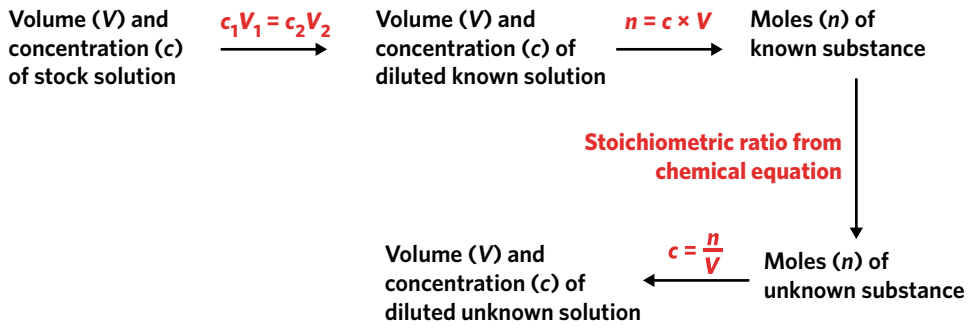


Figure 3 Flowchart of the processes and equations in volume-volume stoichiometry.

One complicating factor occurs when the unknown solution is diluted before performing the titration experiment. This will mean that to determine the concentration of the undiluted known solution, a step is added to the flowchart in figure 3, as shown in figure 4.

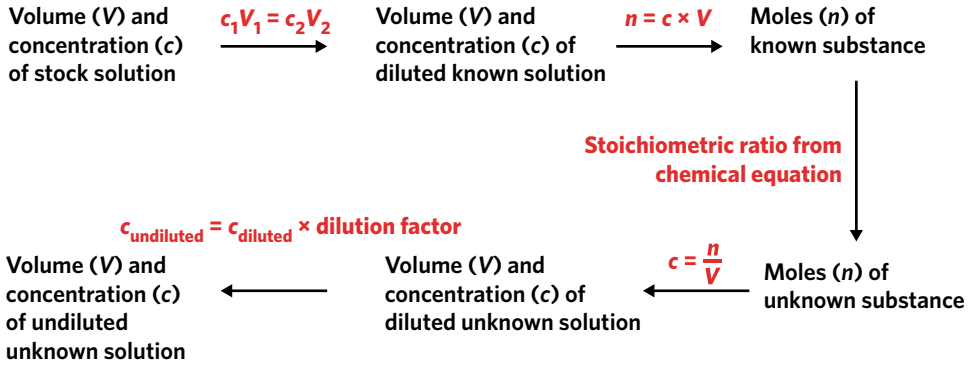


Figure 4 Flowchart of the processes and equation in volume-volume stoichiometry with dilutions involved.

This will allow us to carry out volume-volume stoichiometry calculations because we can link the concentration and volumes of a known solution to the number of moles of substance in that solution. From there, using the stoichiometric mole ratio from the chemical equation, we can calculate the concentration and volume of the unknown solution.

For example, a student has decided to determine the concentration of an unknown solution. She knows the following pieces of information:

Concentration of known solution	2.5 M
Volume of known solution	25.60 mL
Mole ratio from chemical equation (known solution : unknown solution)	1 : 1
Volume of unknown solution	20 mL
Concentration of unknown solution	?

Therefore, as per the flowchart in figure 3, in order to determine the concentration of the unknown solution we first need to calculate the number of moles of known substance.

n(known substance) = $c \times V$ = 2.5 × 0.02560 = 0.064 mol

Next, we need to find the number of moles of unknown substance by using the mole ratio from the chemical equation. In this case,

n(known substance) : n(unknown substance)

1 : 1

n(known substance) = n(unknown substance) = 0.064 mol

Finally, we need to use the volume of the unknown solution to calculate its concentration.

$$c(\text{unknown substance}) = \frac{n}{V} = \frac{0.064}{0.020} = 3.2 \text{ M}$$

Worked example 3

One type of weak acid present in nature is citric acid, $C_6H_8O_7$. It is a triprotic acid and it completely ionises in the presence of a strong base. In order to neutralise a 20.00 mL solution of citric acid, 17.65 mL of 0.20 M NaOH was required. What is the concentration of citric acid present in the solution?

What information is presented in the question?

Citric acid, a triprotic acid, is reacting with NaOH.

Volume of citric acid: 20.00 mL.

Volume of NaOH: 17.65 mL.

Concentration of NaOH: 0.20 M.

What is the question asking us to do?

Find the concentration of the 20.00 mL citric acid solution.

What strategy(ies) do we need in order to answer the question?

1. Determine the chemical equation for the neutralisation reaction between NaOH and $C_6H_8O_7$, and therefore, the mole ratio.
2. Calculate the the number of moles of $OH^-_{(aq)}$ required to neutralise citric acid with the equation $n = c \times V$.
3. Calculate the number moles of citric acid in the 20.00 mL solution using the stoichiometric mole ratio.
4. Calculate the concentration of the citric acid solution with the equation $c = \frac{n}{V}$.

Answer

Since citric acid is a triprotic acid, this means that it is able to donate 3 protons per molecule of citric acid. Therefore, when reacted with sodium hydroxide, NaOH, 1 molecule of citric acid can react with 3 molecules of NaOH. Since this is a neutralisation reaction, an ionic salt and water will be produced as shown by the equation below:

$C_6H_8O_{7(aq)} + 3NaOH_{(aq)} \rightarrow Na_3C_6H_5O_{7(aq)} + 3H_2O_{(l)}$

$n = c \times V$

$n(NaOH) = 0.20 \times 0.01765 = 0.00353$ mol

$n(NaOH) : n(C_6H_8O_7) = 3 : 1$

$$n(C_6H_8O_7) = \frac{1}{3} \times n(NaOH) = 0.0012 \text{ mol}$$

$$c = \frac{n}{V}$$

$$c(C_6H_8O_7) = \frac{0.0012}{\left(\frac{20.00}{1000}\right)} = 0.059 \text{ M}$$

This process of being able to calculate and compare the number of moles of a solution is important in volume-volume stoichiometry as it allows us to find unknown quantities of reactants used or products made. It is also called **volumetric analysis** as we are able to determine the concentration of an unknown solution based on the concentrations and volumes of reacting solutions. These principles will be further investigated later in this lesson.

Principles of titration 2.2.15.2

OVERVIEW

Titration reactions are used to determine the unknown concentration of a solution and the amount of a substance in a solution.

THEORY DETAILS

Sometimes chemists deal with solutions of unknown concentrations. This may be because they are testing a random sample of water from a stream or they want to see if they have prepared a solution correctly. The basic principle of **titration** is that by reacting a solution of known concentration with a solution of unknown concentration, chemists can use their understanding of volume to volume stoichiometry to calculate the concentration of the unknown solution.

As part of this process, a solution of accurately known concentration, also known as the **titrant**, is used to measure the amount or concentration of a substance in a sample solution of unknown concentration, also referred to as the analyte. The basic setup of this reaction can be seen in figure 5.

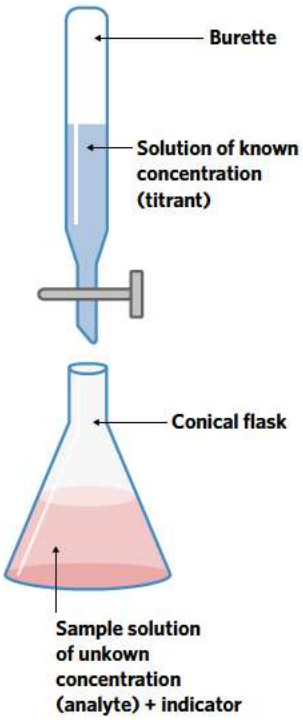
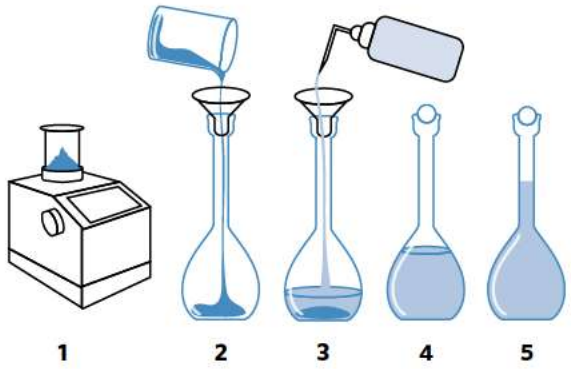


Figure 5 The basic setup for a titration reaction.

Standard solutions

The solution of accurately known concentration usually found in the burette used in titration is known as a standard solution. The titrant used is often standardised (its exact concentration determined) by a solution known as the **primary standard solution**.

To prepare the primary standard solution, a **primary standard** is accurately weighed and dissolved in deionised water in a volumetric flask according to the steps outlined in figure 6. The amount of primary standard used depends on the concentration of the solution we want to prepare.



1. **Mass of primary standard is weighed**
2. **Primary standard is transferred to volumetric flask**
3. **Deionised water added to the volumetric flask until primary standard is dissolved**
4. **Solution is left at rest to allow any volume expansion from newly-formed intermolecular bonds**
5. **Deionised water is added until the meniscus lies on the required volume**

Figure 6 Preparation of a primary standard solution.

The steps required to prepare the primary standard, as illustrated in figure 6, are quite straightforward. Due to the fact that the primary standard solution is used to confirm the concentration of other secondary standard solutions through titration, there are some criteria that need to be met before substances can be considered as a suitable primary standard, such as:

- The substance must be pure and found in a stable form.
- The substance must not react with the atmosphere (e.g. moisture, carbon dioxide etc.).
- The substance must be soluble in solution.
- The substance must be readily available and inexpensive.
- The substance must be able to be weighed accurately (i.e. have a high, known molar mass to reduce error).

An example of a primary standard would be the ionic compound sodium carbonate Na_2CO_3 (anhydrous or 'without water'), which fits all of the above criteria. Common reagents such as NaOH and HCl cannot be used as primary standards as they both will react with the atmosphere. Sodium hydroxide reacts with $CO_{2(g)}$ from the atmosphere to produce sodium carbonate according to the following equation:

$2NaOH_{(s)} + CO_{2(g)} \rightarrow Na_2CO_{3(s)} + H_2O_{(l)}$

This leads to a significant mass increase that will affect any calculations made using the initial concentration of sodium hydroxide. Any concentration of a HCl solution will also change gradually over time, as it has a tendency to vaporise (spontaneously turn into gaseous form) in small amounts when exposed to standard laboratory conditions, again affecting any calculations made using its initial concentration.

Worked example 4

Tom is conducting a titration in order to determine the concentration of ethanoic acid, CH_3COOH, in a sample of water. Outline the steps required to make a 0.0500 M Na_2CO_3 primary standard solution in a 100 mL volumetric flask.

What information is presented in the question?

Volume: 100 mL.

Concentration of Na_2CO_3 = 0.0500 M.

What is the question asking us to do?

Write out the steps needed to prepare a primary standard solution and the quantities involved.

What strategy(ies) do we need in order to answer the question?

1. Calculate the amount, in mol, of $Na_2CO_{3(s)}$ required to be present in the 100 mL volumetric flask with the equation $n = c \times V$.
2. Determine the mass of $Na_2CO_{3(s)}$ required to be weighed on the analytical balance with the equation $m = n \times M$.
3. Write out all the steps required to create this standard solution.

Answer

$n = c \times V$

n (Na_2CO_3) = 0.0500 × 0.100 = 0.00500 mol

$m = n \times M$

m(Na_2CO_3) = 0.00500 × (2 × 23.0 + 12.0 + 3 × 16.0) = 0.0530 g

Therefore, the steps required to produce this primary standard solution are:

1. Weigh 0.530 g of anhydrous sodium carbonate using an analytical balance.
2. Transfer the sodium carbonate to a 100 mL volumetric flask.
3. Add deionised water until sodium carbonate is fully dissolved.
4. Swirl volumetric flask with stopper on and let the solution rest.
5. Add more deionised water until the bottom of the meniscus is at the required mark.

Any factor that may compromise the concentration of the solution will affect the results obtained during the titration reaction. Therefore, we must be careful when choosing a primary standard. Furthermore, we need to be just as careful during the preparation of the standard solution to minimise the chances of errors occurring. For example:

- Incorrect weighing of substances
- Contamination of instruments or substances used
- Rinsing of apparatus with the inappropriate substance

All of these errors can impact the accuracy of the standard solution produced, which will then affect the results obtained.

Titration experiment

To prepare a titration experiment, a **pipette** is used to transfer an accurate volume, called an **aliquot**, of an unknown sample into the **conical flask**. The conical flask is placed under the **burette**, and the burette is then filled with a solution of known concentration. The titrant is dispensed incrementally from the burette into the conical flask, resulting in a chemical reaction in the conical flask. This process is summarised in figure 7.

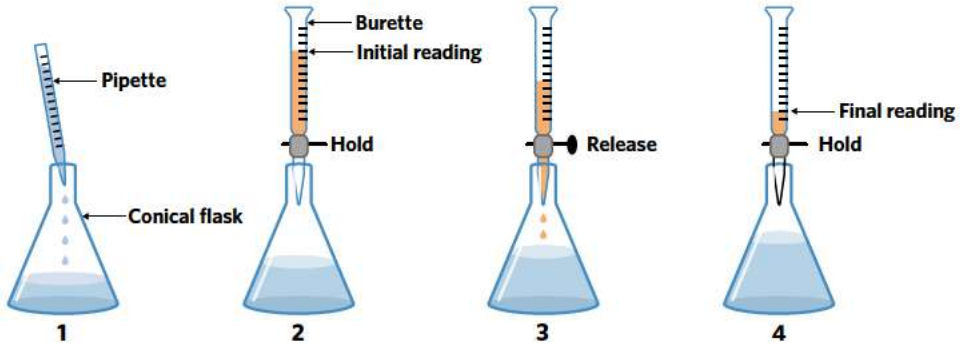


Figure 7 A summary of the titration process.

Depending on the experiment, the reaction will stop at a certain point, at which the appropriate experimental data is recorded and is used to perform calculations that help to determine the concentration of an unknown substance.

It is important to note how to read these values from a burette correctly. Due to the way liquids interact with the walls of the burette, we can usually see a dip in the solution in the burette. Therefore, when we read the scale on a burette, we need to position ourselves so that our line of sight is horizontal to the bottom dip of the **meniscus**. If we don't align ourselves appropriately, the volume read will be higher or lower than the true value as illustrated in figure 8 (depending on whether we are lower or higher than the meniscus), making our results inaccurate.

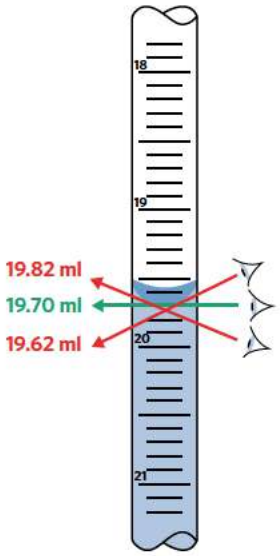


Figure 8 Burette reading.

The experiment is repeated multiple times until **concordant titres** (usually all **titres** within $\pm$ 0.1 mL) are obtained to increase the reliability and precision of the results by minimising the effects of random errors. Table 1 shows an example of the type of data that is recorded during a titration experiment.

Table 1 Sample values from a titration.

Titration number	1	2	3	4	5
Initial burette reading	3.50	7.30	13.35	0.00	5.70
Final burette reading	22.90	26.30	32.40	19.10	23.70
Titre volume (mL)	19.40	19.00	19.05	19.10	18.00

In this case, titration number 2, 3 and 4 are the concordant titres as they are all titres that differ with a 0.1 mL range from the highest value (19.10 mL) to the lowest value (19.00 mL). The **average titre**, which is used in calculations, is the average of the concordant titres.

$$\text{In this case, average titre} = \frac{19.00 + 19.05 + 19.10}{3} = 19.05 \text{ mL}$$

Note that sometimes the initial reading from the burette is lower than the final reading. This is due to the fact that some burettes measure the total volume that has been dispensed and therefore '0.00' is located at the very top of the burette. As such, the more volume that is used during the reaction, the less solution that will be left in the burette, and therefore the larger the reading of the final value, resulting in a larger titre volume.

During titration, the titrant and analyte undergo a chemical reaction. Most of the time, the solutions involved in the reaction are colourless, and therefore we often use an **indicator** to signal that a reaction has taken place. The change in colour that occurs when the indicator is subject to a particular condition (e.g. pH level) signals a particular point in the reaction. This is the point at which the titrant and analyte have reacted to the stoichiometric ratio as indicated in the balanced chemical equation. We can then use this information, and the titre volumes collected to perform appropriate calculations.

Acid-base titrations 2.2.15.3

OVERVIEW

Acid-base titrations are used to determine the concentration of acids and bases via a neutralisation reaction between the two substances.

THEORY DETAILS

As we learned in lesson 8C, acid-base reactions between acids and metal hydroxides are neutralisation reactions in which the hydroxide ions (OH^-) from the base react with hydrogen ions (H^+) from the acid to produce an ionic salt and water. In an acid-base reaction, this hydrogen ion is transferred from the acid to the base. Acid-base titrations are used to determine the concentration of an unknown base or acid.

Building on our knowledge of titration reactions, there are a few key characteristics of acid-base titrations that we need to understand. As discussed in lesson 8B, there are various indicators to help signal a change in pH values during a reaction involving an acid and a base. We can incorporate the colour change experienced by the indicator in a titration reaction to help signal a particular point of the acid-base reaction.

When preparing an acid-base titration, we need to add an appropriate indicator to the analyte solution in the conical flask prior to dispensing the titrant into the flask (and thereby starting the acid-base reaction). After the addition of a particular volume of the titrant, the solution in the conical flask will change colour. The acid-base titration reaction is summarised in figure 9.

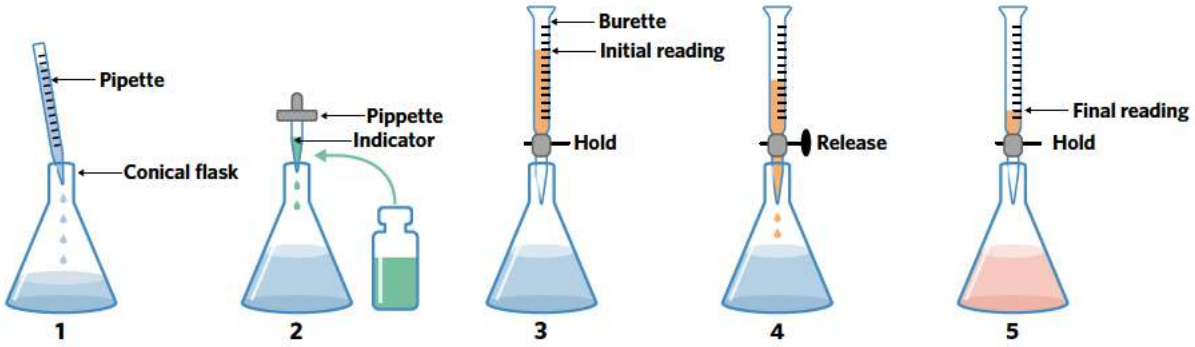


Figure 9 Acid-base titration.

The colour change that occurs represents the **end point** of the titration. This indicates the end of the titration, and that the titrant and analyte have reacted to the stoichiometric ratio as indicated in the balanced chemical equation, referred to as the **equivalence point**. Due to this, it is very important that we choose an appropriate indicator in which the colour change occurs within the range of the equivalence point. It is important to note that the end point generally occurs just before the equivalence point. However, for calculation purposes, we assume that the equivalence point is approximately the same as the end point.

The volume of titrant, known as the titre, used to reach the end point is measured and used to calculate the amount or concentration of the analyte solution. This volume is calculated by noting the initial volume reading on the burette as well as the final burette reading. The difference between these two values is the volume of solution delivered. After multiple repetitions of the titration , the average value of concordant titres, (i.e. average titre), obtained during the reaction is used in calculations to determine the unknown value.

Calculations involving acid-base titrations

The values obtained from a titration reaction can be used to quantitatively determine unknown concentration of any of the substances involved in the reaction.

Consider the reaction between nitric acid (HNO_3) and potassium hydroxide (KOH):

$$HNO_{3(aq)} + KOH_{(aq)} \rightarrow KNO_{3(aq)} + H_2O_{(l)}$$

Suppose a student wanted to find the concentration of an unlabeled bottle of potassium hydroxide. As a result of a titration reaction, it was shown that an average titre of 24.12 mL of a 0.50 M sample of nitric acid was used to neutralise a 20.00 mL aliquot of potassium hydroxide. How could the student use this information to determine the unknown concentration of potassium hydroxide?

Considering the data collected from the experiment represents the end point of the reaction, and therefore if the right choice of indicator was used, equivalence point, we need to first determine the amount of the known substance, nitric acid, used in the reaction:

$$n(HNO_3) = c \times V = 0.50 \times \left(\frac{24.12}{1000}\right)$$

$$= 0.01206 \text{ mol}$$

We can see that a total of 0.01206 mol of nitric acid was required in the reaction. According to the balanced equation for the reaction,

$n(HNO_3) : n(KOH)$

1 : 1

Therefore, for every 0.01206 mol of nitric acid used in the reaction, a total of 0.01206 mol of potassium hydroxide would also be required. Since a 20.00 mL aliquot of $KOH_{(aq)}$ was used, we can now use that information to determine the unknown concentration of the potassium hydroxide solution:

$$c(KOH) = \frac{0.01206}{\left(\frac{20.00}{1000}\right)}$$

= 0.60 M

An overall summary of this process can be seen in figure 10.

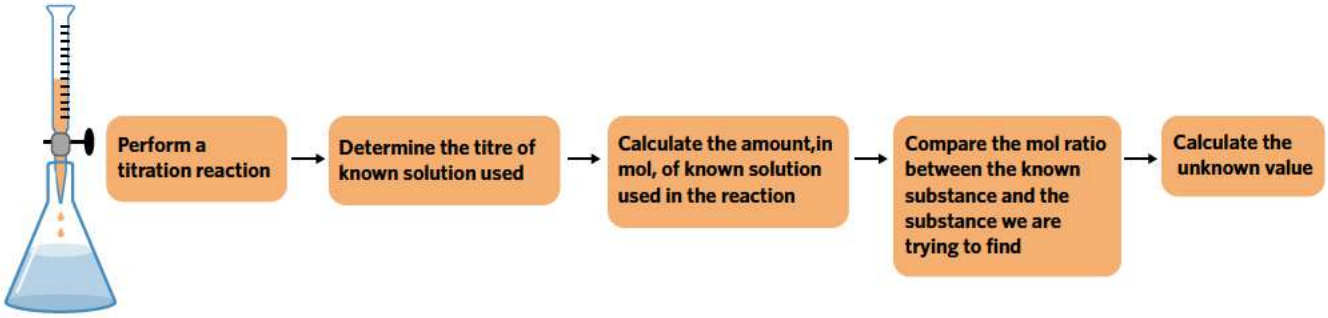


Figure 10 Process involved in calculating the value of an unknown solution using titration .

Worked example 5

A chemistry student was asked to determine the concentration of a sample of the basic magnesium hydroxide solution, $Mg(OH)_2$. After conducting multiple titrations with a 25.00 mL aliquot of magnesium hydroxide, she found that it took an average of 25.60 mL of a 0.50 M hydrochloric acid, HCl, solution for a colour change to occur. Calculate the concentration of the magnesium hydroxide solution.

What information is presented in the question?

25.00 mL aliquot of magnesium hydroxide.

25.60 mL of 0.50 M hydrochloric used.

What is the question asking us to do?

Find the concentration of the magnesium hydroxide solution.

What strategy(ies) do we need in order to answer the question?

1. Determine the amount of hydrochloric acid used in the reaction with the equation $n = c \times V$.
2. Use stoichiometry to determine the mole ratio of HCl to $Mg(OH)_2$.
3. Determine the amount of magnesium hydroxide present.
4. Determine the concentration of the magnesium hydroxide solution with the equation $c = \frac{n}{V}$.

Answer

The balanced equation for this reaction is

$2HCl_{(aq)} + Mg(OH)_{2(aq)} \rightarrow MgCl_{2(aq)} + 2H_2O_{(l)}$

$n(HCl) = 0.50 \times 0.02560$

$= 0.013$ mol

$n(HCl) : n(Mg(OH)_2)$

2 : 1

$$n(Mg(OH)_2) = 0.013 \text{ mol} \times \frac{1}{2}$$

$= 0.0064$ mol

$$c(Mg(OH)_2) = \frac{0.0064}{\left(\frac{25.00}{1000}\right)}$$

$= 0.256$ M

$= 0.26$ M

Representing acid-base titrations

Acid-base reactions can be represented through a titration curve which shows the change in pH of the analyte solution as titrant is added from the burette. As we know, the change in pH throughout the reaction depends on the concentration of hydroxide and hydrogen (or hydronium H_3O^+) ions found in the analyte solution.

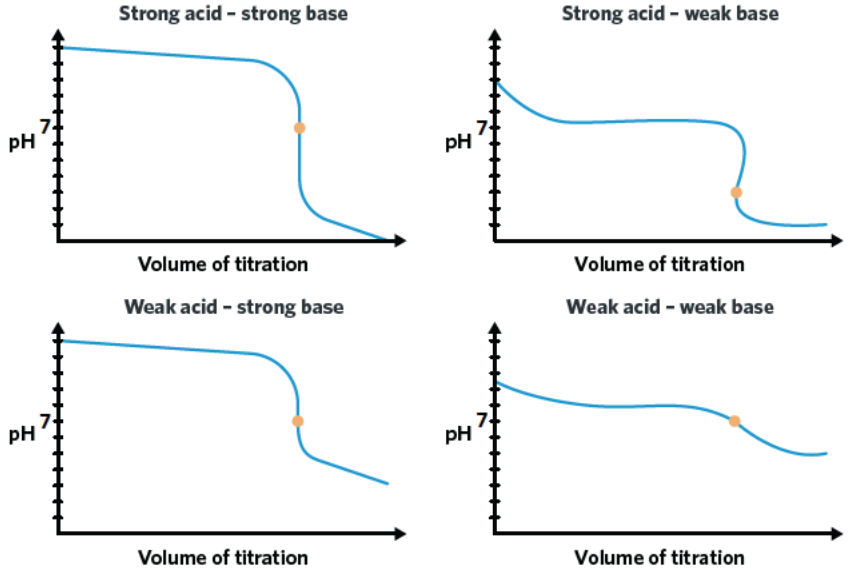


Figure 11 Example of a titration curve involving the titration of different strengths of acids and bases.

As shown in figure 11, the equivalence point can be located at the point during which the curvature of the curve starts to change. This can sometimes be referred to as the point of inflection. The starting point and shape of the titration curves can help us to determine the nature of the analyte. For example in figure 11, we can see that the analyte began as a basic solution as the curve started at a high pH and progressively became more acidic. From this, we can imply that the analyte solution is basic and the titrant is acidic. If the analyte solution were acidic and the titrant basic, the curve would be going from a low pH to a high pH.

It is important to know the approximate pH of the equivalence point. This is because when choosing an appropriate acid-base indicator, we need to make sure that the equivalence point falls within the pH range for the colour change of the indicator, which is the observed end point of the titration. For example, when a strong acid reacts with a weak base, we would expect that the pH of the solution at the equivalence point would be lower than 7 (i.e. more acidic) and therefore need to pick an acid-base indicator that will change colour within this range. Figure 12 displays the list of indicators in the Chemistry Units 3 & 4 data book, their colour changes and what pH range they change in.

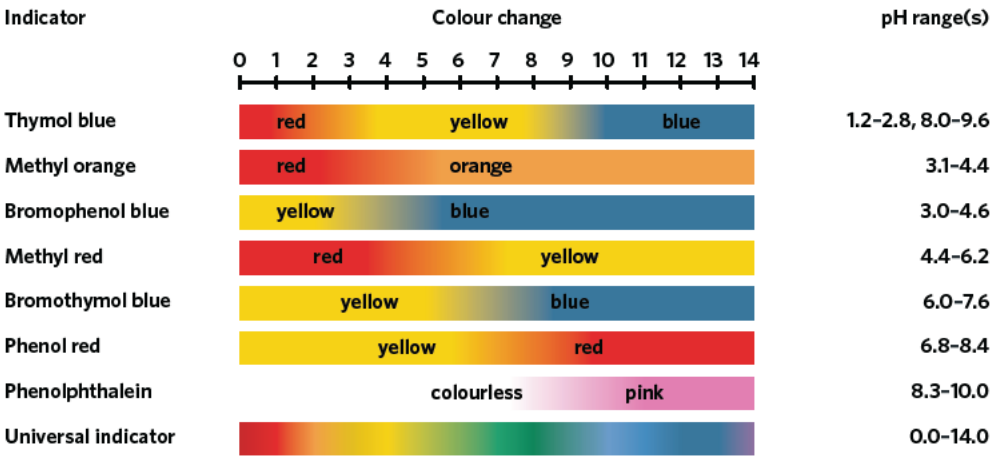


Figure 12 List of indicators used in titrations.

Consider the reaction between a strong acid and a weak base. Since the equivalence point will occur at a pH lower than 7, bromophenol blue could be an appropriate choice of indicator. This is because it changes from blue to yellow in the 3.0–4.6 pH range.

Therefore, as shown in figure 13, the person performing the titration could accurately determine the equivalence point of the titration from the end point signalled by the colour change.

It is important to note that the equivalence point should lie within the pH range of the indicator, rather than at the ends of its pH range. This is to ensure that a complete colour change occurs as close as possible to the equivalence point, rather than too early or too late.

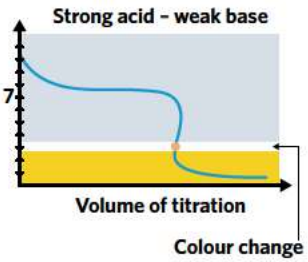


Figure 13 Titration curve of a chemical reaction involving a strong acid and weak base with bromophenol blue.

Experimental errors

Due to the practical nature of titration, experimental errors can be introduced in a variety of ways. These could include misreading values on the burette or using the wrong choice of indicator. Mistakes in experimental procedure tend to arise from the incorrect washing of glassware. The burette and pipette have to be washed with the acid/base that they are filled with. This is because if distilled water is used, the acid/base will be diluted, resulting in a lower concentration of acid/base being delivered. This will then impact the accuracy of the results. On the other hand, the volumetric flask has to be washed with distilled water. This is because if it is rinsed with acid/base then extra amounts will be added which were not calculated, affecting the concentration of the analyte under investigation in the conical flask. These mistakes are summarised in figure 14.

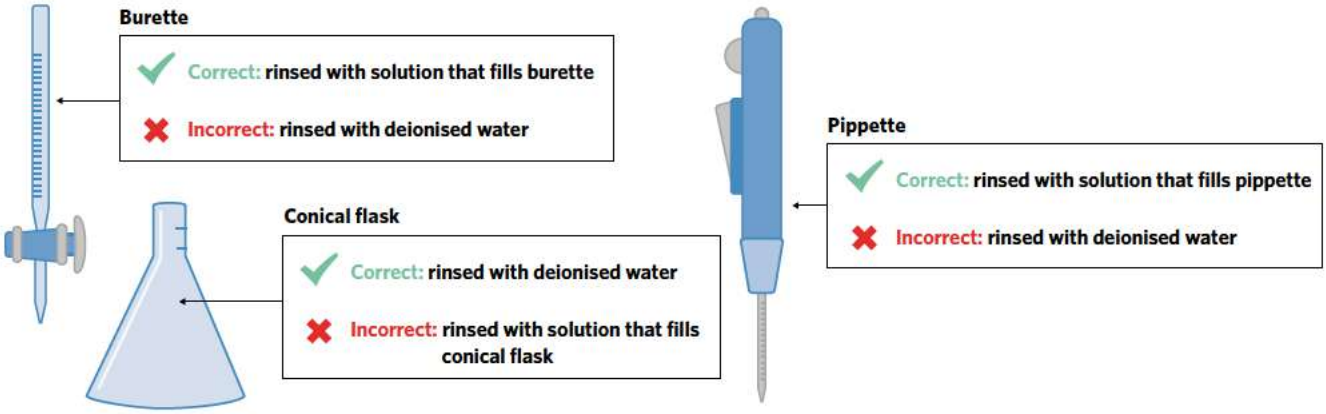


Figure 14 Mistakes arising from the incorrecting rinsing of titration glassware.

Theory summary

- Titration is a quantitative technique to determine the unknown concentration of a solution and/or the amount of a compound in a solution.
- Titrations use standard solutions of known concentration to determine the concentration of unknown solutions.
- The equivalence point is where the exact stoichiometric ratio of acid to base is present, whereas the end point is when the indicator changes colour.
- Primary standards need to be readily available, pure and not react with components in the air as it may compromise the concentration of the standard solution.
- Titration curves of acid-base reactions show pH changes and can be used to determine the appropriate choice of indicator.

10F QUESTIONS

THEORY REVIEW QUESTIONS

Question 1

A primary standard must be

A unreactive with the atmosphere.

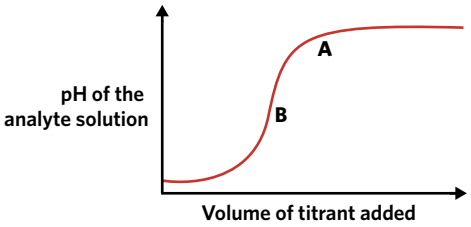
B insoluble.

Question 2

Which point on the graph represents the equivalence point?

A Point A

B Point B



Question 3

The equivalence point is the same as the end point in every titration.

A True

B False

Question 4

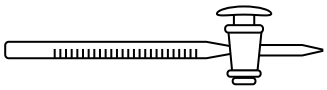
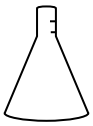
When the indicator changes colour, this signals that the _____ has been reached.

A equivalence point

B end point

Question 5

Which option correctly matches the type of glassware used in titration with its function?

	Glassware	Function
A		Dispenser of the analyte
B		Vessel for the acid-base reaction

Question 6

Which of the following options contains exactly three concordant titres?

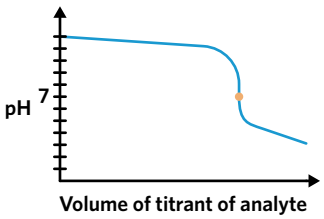
	Recorded titres
A	17.65, 17.62, 18.05, 17.64, 17.60
B	17.65, 17.72, 18.05, 17.64, 17.60

Question 7

Based on the titration curve shown, which of the following indicators would be most appropriate for this acid-base titration?

A Methyl red

B Phenol red



Question 8

When reading the burette shown below, the value recorded is the _____ of the meniscus.



A top

B bottom

Question 9

Common strong acids and bases like HCl and NaOH make excellent standard solutions.

A True

B False

Question 10

In order to make a standard solution, a primary standard is weighed on an analytical balance, which is a very accurate measurement device. Why is it important to use an analytical balance?

A The mass of primary standard used is important in determining the concentration of the standard solution.

B The mass of primary standard used has to always be the same in standard solutions.

Deconstructed questions

Use the following information to answer questions 11–13.

Fabian, an enthusiastic chemistry student, decided to set up a titration experiment for his practical report. He is investigating the concentration of sodium hydroxide in different brands of household cleaners. He decides to use 1.0 M hydrochloric acid as the titrant in order to neutralise the sodium hydroxide present in a 20.00 mL aliquot of household cleaner. His table of results for one cleaner, SutherClean^{TM}, is recorded below:

Titration number	1	2	3	4	5
Final burette reading	22.50	23.75	32.60	34.90	27.80
Initial burette reading	2.50	3.70	13.00	14.95	8.30
Titre volume (mL)	20.00	20.05	19.60	19.95	19.50

Question 11 (1 MARK)

What is the neutralisation equation occurring in the conical flask during the titration?

A $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

B $HCl_{(aq)} + NaCl_{(aq)} \rightarrow NaOH_{(aq)} + H_2O_{(l)}$

C $NaCl_{(aq)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + HCl_{(aq)}$

D $Cl^-_{(aq)} + Na^+_{(aq)} \rightarrow NaCl_{(aq)}$

Question 12 (1 MARK)

Which of the following titration numbers produced concordant results?

A 1, 2, 3

B 2, 3, 4

C 2, 4, 5

D 1, 2, 4

Question 13 (3 MARKS)

Calculate the concentration of sodium hydroxide present in SutherClean^{TM}.

Exam-style questions

Within lesson

Question 14 (9 MARKS)

The Japanese giant hornet, *Vespa mandarinia japonica*, is the largest wasp in the world. Chris, having recently been stung by one of these wasps, decided to investigate the properties of its sting. He knows that present in this wasp's sting is magnesium hydroxide, $Mg(OH)_2$ and carries out an acid-base titration to determine the concentration of magnesium hydroxide in the sting in order to design a therapeutic treatment that precisely neutralises the sting using an acid. He knows that each wasp sting delivers 0.30 mL of fluid. His choice of titrant, 0.0010 M HCl, reacts with $Mg(OH)_2$ in a 2 : 1 ratio. For the titration, he uses a 20.00 mL aliquot of wasp sting fluid.

a How many wasp stings does Chris need to extract the venom from to carry out this titration? (1 MARK)

b He uses a special indicator for this titration. What is the purpose of the indicator in a titration? (1 MARK)

His table of results is shown:

Titration number	1	2	3
Final burette reading	24.60	29.50	23.60
Initial burette reading	2.50	7.35	1.50
Titre volume (mL)	22.10	22.15	22.10

c Are these concordant titres? Explain why or why not. (2 MARKS)

d Calculate the concentration of magnesium hydroxide present in the aliquot. (3 MARKS)

e Calculate the amount of magnesium hydroxide present in each wasp sting. Based on this, what volume of 0.0010 M HCl should be applied to the wasp sting for treatment? (2 MARKS)

Question 15 (6 MARKS)

Biodiesel is a viscous fuel that you will study in more detail for VCE Chemistry Units 3 & 4. Biodiesel can be produced from waste vegetable oils. It is important for biodiesel producers to determine the pH of waste vegetable oils which are slightly acidic, however, normal techniques of measuring pH like litmus paper or a pH probe are not able to be used.

a Suggest a reason why titration is used to determine the pH of the vegetable oils rather than the traditional methods. (1 MARK)

An excerpt of the experimental method is shown below to determine the pH of biodiesel:

1 Add 1 mL of waste vegetable oil to the conical flask, accompanied by 10 mL ofpropanol.

2 Add a few drops of the bromothymol blue indicator.

3 While swirling the conical flask, add 1.0 M NaOH from the burette drop by drop until the indicator permanently transitions from yellow to blue colour.

b With respect to this experimental method, explain the difference between the equivalence point and the end point. (3 MARKS)

c The analyst performing this experiment did not use a primary standard solution to standardise the NaOH titrant used in the experiment. Why is the use of a primary standard solution important? (2 MARKS)

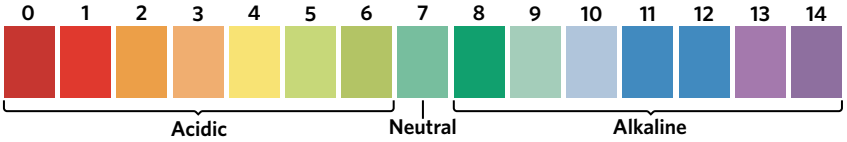
Question 16 (8 MARKS)

Acid-base indicators are used to determine the point a solution is acidic, basic or neutral through a colour change. This colour change is caused by the fact that most indicators are actually weak acids and bases themselves and actually react with the acid and base in the titration. Therefore, when they react, the conversion to the conjugate acid/base of the indicator is what causes the colour change.

a Given that acid-base indicators are weak acids and bases themselves, would adding more indicator increase or decrease the accuracy of the experiment? Explain why or why not. (3 MARKS)

b For a given titration, the equivalence point is at a pH of approximately 6.8. What would be the best choice of indicator to use? Justify your answer with reference to the definitions of the equivalence and end points. (3 MARKS)

c Universal indicator is an acid-base indicator that has a colour spectrum that ranges from very acidic conditions to very basic conditions.



It contains thymol blue, methyl orange, methyl red, bromothymol blue and phenolphthalein. Describe a possible reason for how the rainbow of colours above could be created from just five indicators. (2 MARKS)

Question 17 (8 MARKS)

Sodium carbonate, Na_2CO_3 is a chemical compound that is commonly used as a primary standard. It is commercially available in a form that has some water present (i.e. it is not completely dry).

a In order for sodium carbonate to be effective as a primary standard, what needs to happen to it between its purchase from a commercial source and its use. Explain your answer. (2 MARKS)

b List three properties that define what makes a good primary standard. (1 MARK)

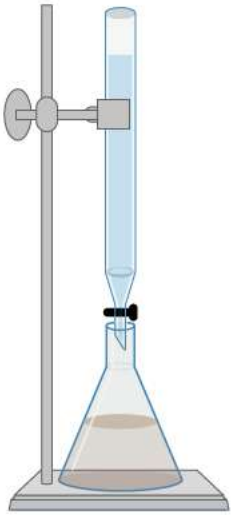
c In the preparation of a sodium carbonate primary solution, a volumetric flask is used. Why is a volumetric flask used as opposed to other types of glassware? (1 MARK)

d Potassium hydrogen phthalate (M_r = 204.2 g mol^{-1}) is commonly used as a primary standard to standardise sodium hydroxide solutions. If 250.00 mL of a 1.000 M potassium hydrogen phthalate solution is required, outline the steps involved in the preparation of this standard solution, including all relevant calculations. (4 MARKS)

Multiple lessons

Question 18 (9 MARKS)

Food is carefully monitored and regulated in Australia to ensure that products are safe for human consumption. Therefore, vinegar and its acidity is carefully controlled. The acidity of vinegar is caused by acetic acid, CH_3COOH. Vinegar acidity is closely monitored because of its importance in food preservation where a certain %(v/v) of acetic acid is required to kill harmful microbes and bacteria.



The Food Safety Board decided to investigate the pH of a new brand of vinegar by titration with the following procedure:

1. 25.00 mL aliquot of vinegar was pipetted into a conical flask.
2. 3 drops of phenolphthalein indicator were added to the vinegar.
3. The burette was filled with standardised 1.0 M NaOH.
4. NaOH was dripped into the conical flask until the indicator turned from colourless to pink.
5. The burette values were recorded and the experiment repeated until concordant titres were obtained.

The results are shown below:

Titration number	1	2	3	4	5
Final burette reading	21.62	21.43	21.58	21.84	21.61

a Write the full balanced equation for the reaction between NaOH and acetic acid. (1 MARK)

b What is the name of the functional group that gives acetic acid its acidic properties? (1 MARK)

c Calculate the average titre volume. (2 MARKS)

d Calculate the concentration of acetic acid present in the vinegar. (2 MARKS)

Acetic acid has a density of 1.05 g mL^{-1} at 25°C. It is regulated that the %(v/v) of acetic acid has to be greater than 4%.

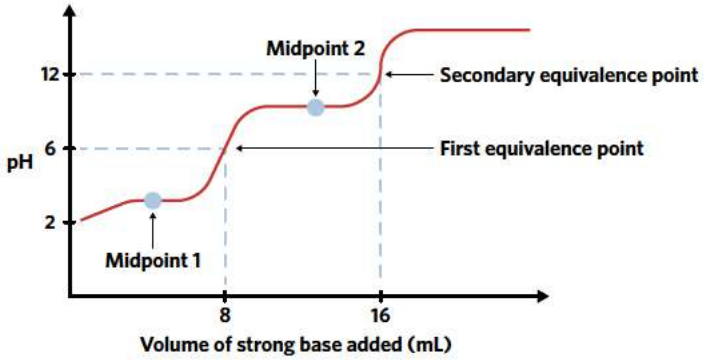
e Calculate the %(v/v) of acetic acid in the vinegar. Does this new brand of vinegar meet the regulatory standards? (3 MARKS)

Question 19 (8 MARKS)

There are many different types of acids, each with different strengths and properties.

a What is a polyprotic acid? (1 MARK)

The titration curve for a diprotic acid is shown below:



b Why are there two equivalence points? (2 MARKS)

c What choice of indicator(s) will be suitable to identify both equivalence points? (2 MARKS)

When performing this titration, the concentration of acid in the analyte was too large for the base, and therefore, was diluted. The solution was diluted by a factor of 10.

d If this acid-base titration had a stoichiometric ratio of 1 : 1 of acid to base, and 16.00 mL of a 1.0 M basic solution was used to neutralise a 20.00 mL aliquot of the solution of acid, after the above dilution, what is the concentration of acid in the undiluted solution? (3 MARKS)

Key science skills

Question 20 (8 MARKS)

It is important in titrations to wash the glassware before using it. This is because the glassware may not have been washed correctly or at all before being packed away. In one titration, a student is titrating 1.0 M HCl with a 20.0 mL aliquot of NaOH solution of unknown concentration.

a If the burette is washed with water, what effect would this have on the accuracy of the calculated concentration of NaOH solution? Justify your answer. (3 MARKS)

b Why is the conical flask washed with distilled water rather than the NaOH solution? (2 MARKS)

c Glassware is fragile and can be easily broken. Unfortunately, a student accidentally knocks over a retort stand and the burette, along with the acid inside, is knocked on the floor and smashes. Suggest three safety precautions that have to be taken to minimise the effect of this incident. (3 MARKS)

Exam-style question hints

14a: The number of wasps required depends on the volume of venom required and how much venom each wasp sting delivers. **14b:** Since most acids and bases are colourless, the experimenter needs to know when to stop adding titrant to the analyte solution. **14c:** Concordant titres differ from each other by a certain value. **14d:** Titrations determine the number of moles of the unknown compound and then the concentration of the unknown solution. **14e:** The aliquot of $Mg(OH)_2$ contains many wasp stings. **15a:** pH measurements usually take place in aqueous solutions. **15b:** The end point and the equivalence point are different. **15c:** NaOH is not a primary standard.**16a:** Only a few drops of indicator are added in a titration. **16b:** Indicators indicate the end point of a reaction. **16c:** There are many different types of indicators. **17a:** Water adds to the weight of substances. **17b:** Primary standards have specific required properties to perform their function. **17c:** Volumetric flasks are highly accurate pieces of glassware. **17d:** Primary standards solutions are prepared to ensure solutions of an accurately known concentration. **18a:** Acid-base reactions with metal hydroxides produce water and an ionic salt. **18b:** Functional groups are special arrangements or groups of atoms that give compounds certain chemical and physical properties. **18c:** To calculate the average titre, the concordant titres need to be determined. **18d:** The stoichiometric ratio of acid to base according to the chemical equation is important in titration. **18e:** The relates to the volume of acetic acid per volume of vinegar solution. **19a:** 'poly' is a prefix that means 'many'. **19b:** Equivalence points relate to when an acid is fully deprotonated. **19c:** Different indicators are used depending on the pH of the equivalence point. **19d:** Dilution factors relate the concentration of the diluted solution to the undiluted solution. **20a:** The concentration of titrant delivered has to be accurate in order to calculate the unknown concentration of analyte correctly. **20b:** The amount of analyte in the volumetric flask is what the calculation is based upon, not the concentration of analyte as this is unknown. **20c:** The type of precautions taken depends on the nature of the safety hazard.

10G CHROMATOGRAPHY

In this lesson we will be learning about the principles of chromatography and how it is used to quantitatively and qualitatively analyse samples.

10A Water on Earth | 10B Water contamination | 10C Analysing water using gravimetric analysis | 10D Analysing water using light - part 1 | 10E Analysing water using light - part 2 | 10F Volumetric analysis | 10G Chromatography

Study design dot point

- the application of high performance liquid chromatography (HPLC) including the use of a calibration curve and retention time to determine the concentration of a soluble organic compound in a water sample (excluding details of instrument)

Key knowledge units

Principles of chromatography	2.2.13.1
HPLC	2.2.13.2
Analysing chromatograms	2.2.13.3

Key terms and definitions

Chromatography analytical technique used to determine the identity of components in a mixture as well as their concentration

Chromatogram visual output of chromatography which can be analysed

Stationary phase solid onto which the components of a sample adsorb

Mobile phase phase that flows through the stationary phase carrying components of a sample with it

Adsorption (chromatography) component adheres to the stationary phase

Desorption a component is released from the stationary phase and dissolves into the mobile phase (in the context of chromatography)

HPLC (High performance liquid chromatography) a type of chromatography technique which pumps the mobile phase through a very tightly packed stationary phase under high pressure

Retention time, R_t measure of time taken for a component to travel through a chromatography column eg HPLC column

Elute (used for HPLC) movement of a substance out of the chromatography column

Eluent combination of the mobile phase and sample that moves through the HPLC column

Lesson links

This lesson builds on:

- 5C Intermolecular bonding

Intermolecular bonds are important in the separation of components in chromatography.

Principles of chromatography 2.2.13.1

OVERVIEW

Chromatography is a common analytical tool used to identify unknown components of a sample based upon the unique retention time of each component.

THEORY DETAILS

Chromatography is a widely used technique for separating and analysing complex mixtures of organic compounds. This technique uses the properties of different compounds as a way to separate substances from a larger mixture.

All methods of chromatography have a **stationary phase** and a **mobile phase**. These phases will interact differently with different substances depending on their structures. The mobile phase is what the components of a sample dissolve into and is, for the most part, a liquid. The stationary phase is the surface that the mobile phase (and subsequently sample) flows over. This technique relies on the following two key ideas:

1. How well the sample sticks to the stationary phase through a force of attraction (this is known as **adsorption** to the stationary phase)
2. How well the mobile phase can dissolve the sample through an attractive force (this is known as **desorption** into the mobile phase)

Let's use an analogy to help us understand chromatography. A swarm of insects is approaching some flowers. The bees are much more attracted to the flowers than the wasps. As a result, the bees spend longer with the flowers compared to the wasps. This leads to separation as shown in figure 1. In this case, the flowers represent the stationary phase and the wind pushing the insects across the flowers is the mobile phase.

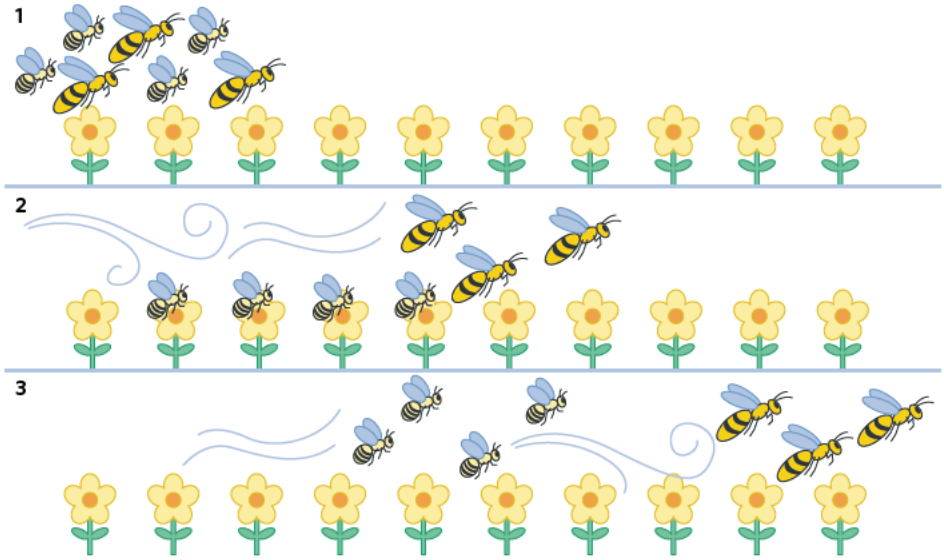


Figure 1 An analogy for chromatography.

One of the main ways in which compounds are separated in chromatography is based on polarity. As we learned in lesson 5B, compounds can be classified as being polar or non-polar depending on their bonding and three-dimensional structure. Given that non-polar compounds interact with other non-polar compounds and polar compounds interact with other polar compounds, we can use this property to separate organic compounds from mixtures

For example, if we have a polar stationary phase, we know that it is capable of permanent dipole-dipole bonding. This means that if a component in the sample is also polar it will be strongly adsorbed to the stationary phase. However, if the mobile phase is polar instead, a polar component will be more strongly desorbed into the mobile phase. Since the components of a mixture have different structures, they will all have different types of intermolecular interactions with varying strength. In turn, the components will be attracted to the stationary phase and the mobile phase differently. As a consequence, each unique substance will have its own specific **retention time** and we can easily recognise each one from the time it takes to **elute** or come out of the chromatography column.

The reason this technique is commonly used is due to its ability to extract components of a mixture without irreparably damaging the sample. As a result, the samples can be used for further analysis if required.

HPLC 2.2.13.2

OVERVIEW

High performance liquid chromatography (**HPLC**) is a chromatography technique that consists of a solid stationary phase which is tightly packed into a column, usually glass, and a solvent acting as the mobile phase.

THEORY DETAILS

High performance liquid chromatography (HPLC) is a highly sensitive analysis technique which is widely used in laboratories across the world. For example, HPLC can be used by sporting authorities to detect drug-related compounds in athletes.

In order to achieve the greatest separation possible, a column is set up with a solid stationary phase which consists of tiny particles as shown in figure 2. Having small particles increases the surface area of the column, which therefore allows for a large number of interactions between the stationary phase and the components in the sample.

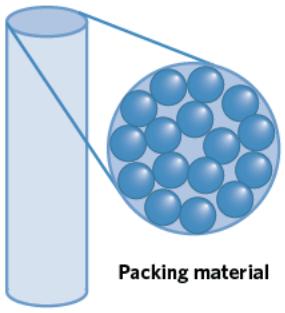


Figure 2 The composition of a HPLC column.

However, the smaller the particles, the harder it is for the particles to flow. As a result, the mobile phase has to be pumped in at high pressure; hence the name of this technique. A diagram of a column used in HPLC is shown in figure 3.

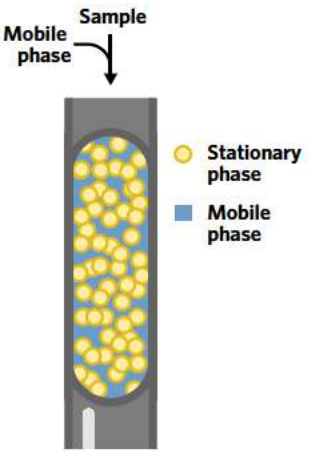


Figure 3 Diagram of a section of the column used in HPLC.

A component that is more strongly adsorbed to the stationary phase will spend more time in this phase compared to another component which is less strongly adsorbed to the stationary phase. Likewise, a component that is more soluble in the mobile phase will spend more time in the mobile phase compared to another component which is less soluble in the mobile phase. In simpler terms, the components experience different rates of adsorption to the stationary phase and desorption off the stationary phase and therefore will leave the column, or will elute, at different times, leading to different retention times. Retention time in this context is the time that a component of a sample spends in the HPLC column. The full setup of this technique is illustrated in figure 4.

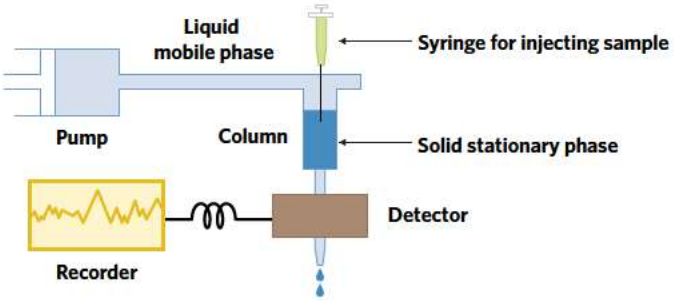


Figure 4 Diagram of the elements of HPLC.

Since many organic compounds are clear and cannot be seen with the eye when emerging from the end of the column, a recorder is used to detect the presence of organic compounds based on their absorption of UV light. Therefore, if an organic compound is present in the **eluent** stream (what comes out of the chromatography column), the recorder will detect a decrease in the UV light signal. This is recorded on a chromatogram.

The retention time of each component depends on the extent to which the components in the sample are attracted to the stationary/mobile phase. This is used to identify the components of a sample because, if carried out under the same conditions, a component will always have the same retention time and therefore, retention time acts as a way to confirm the presence of a specific component in the sample.

Due to the importance of the forces of attraction between each component and the stationary and mobile phase, there are a variety of factors which can influence retention time. These include:

- Identity and composition of the stationary phase
- Identity and composition of the mobile phase
- Length of the column
- Temperature of the column
- Mobile phase flow rate
- Total surface area of the stationary phase

The effect of these changes are outlined in table 1.

Table 1 The effect of a change to conditions on retention time.

Change to conditions	Effect on retention time	Description
Longer column	Increased	There are more particles that components need to move through within the column and therefore, it will take a longer time to elute.
Increased temperature	Decreased	Particles have more energy to move, and therefore can move through the column a lot faster.
Increased flow rate of mobile phase	Decreased	Particles are being 'swept away' faster in the mobile phase and therefore move through the column a lot faster.
Increased total surface area of stationary phase	Increased	There are more interactions between the component and the stationary phase and therefore, it will take longer to elute.

All of these changes in some way affect the interactions that occur between components of a sample and the stationary and mobile phase. For example, when the temperature increases, so does the solubility of most substances, excluding gases. This means that in HPLC, an increase in temperature (and thereby solubility) will result in components being more strongly desorbed into the mobile phase, leading to shorter retention times.

Analysing chromatograms 2.2.13.3

OVERVIEW

Chromatograms produced by HPLC can reveal important information about the composition of a mixture. This information can either be qualitative (i.e. the identities of components in a mixture) or quantitative (i.e. how much of a component there is).

THEORY DETAILS

Without calibration, chromatography is purely qualitative as it allows us to find the identities of different components based on their retention times. Suppose a clinician wanted to measure the levels of dopamine, an important neurotransmitter linked to Parkinson's Disease, in certain areas of the body. Upon running a sample through a HPLC, a chromatogram was produced as shown in figure 5.

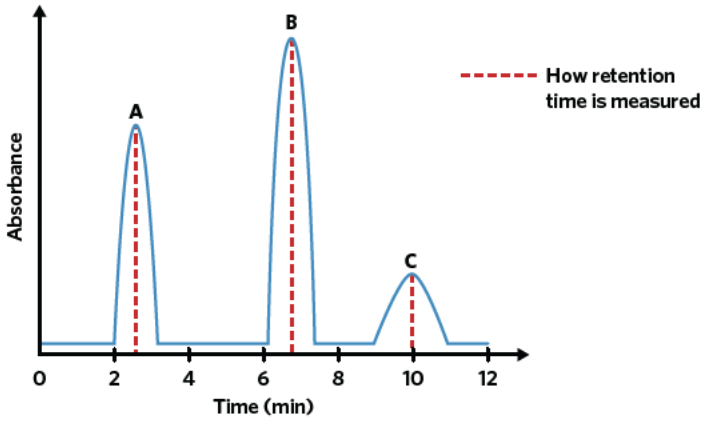


Figure 5 A sample HPLC chromatogram.

Each peak in the chromatogram represents a separate component in the sample. The three peaks indicate that there are three components in this sample and the retention time of each component is measured from the tip of each peak. The clinician also knows that under the exact same conditions the following retention times were recorded.

Table 2 Components and their respective retention times.

Identity of component	Retention time (min)
Dopamine	10
Norepinephrine	2.5
Epinephrine	6.6

Therefore, by matching the observed retention times to the retention times of known samples, the clinician can qualitatively determine the identity of each of the peaks.

The other powerful piece of information that HPLC can provide relates to quantitative data, which tells us 'how much' of a substance is present. The amount of a substance is proportional to the area under its peak and absorbance. Therefore, a chromatogram can also reveal the amount of a substance present. To do this, a set of standards of known concentrations needs to be tested in the same HPLC column under the exact same conditions as the experiment itself. Using the results, we can draw a calibration curve like the one shown in figure 6 to show the relationship between the area under the peak and the concentration or amount of substance present.

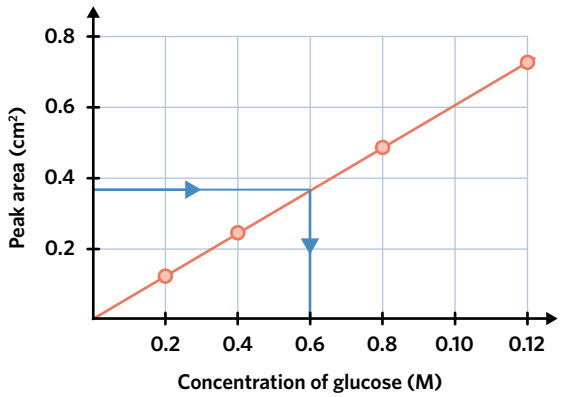


Figure 6 Calibration curve comparing peak area and the concentration of glucose.

In this case, a clinician's unknown sample had a peak area of around 0.36 cm. Therefore, by drawing a calibration curve using a 'line of best fit', this value can be matched to an appropriate concentration which is 0.6 mg/L. It is also important to note that a calibration curve is only accurate for concentrations that are contained within the upper and lower bounds of the samples. As a general rule of thumb, we also want to have at least two data points both above and below our unknown concentration to ensure a high level of accuracy. If we created a calibration curve using concentrations that did not include the peak area we wanted to investigate, we would not be able to accurately determine the concentration of the substance.

Theory summary

- Chromatography is an analytical tool used to determine the components of a sample as well as their concentration.
- The output of a chromatography instrument yields a chromatogram, which can be used to identify components of a sample.
- A chromatogram produced by HPLC shows the retention times of the components which can be matched to those of known organic compounds calculated under the same conditions.
- A calibration curve can be constructed by running samples of a known concentration which can relate the area under the peak to the concentration of the component in the sample.

10G QUESTIONS

Theory review questions

Question 1

Chromatography can only be used to separate components of a mixture.

A True

B False

Question 2

In the process of chromatography,

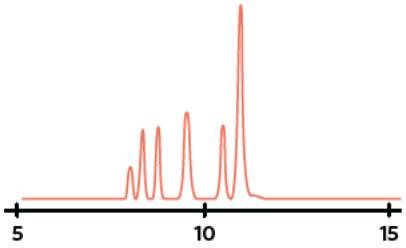
A components are desorbed into the mobile phase once and adsorbed to the stationary phase once.

B continuously desorbing into the mobile phase and adsorbing to the stationary phase.

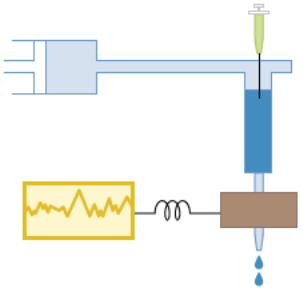
Question 3

Which of the following images represents a chromatogram?

A



B



Question 4

In HPLC,

A the solvent (liquid) is the stationary phase and the solid in the glass column is the mobile phase.

B the solvent (liquid) is the mobile phase and the solid in the glass column is the stationary phase.

Question 5

Retention time depends solely on the identity of the stationary and mobile phase.

A True

B False

Question 6

If a polar stationary phase is being used, components which are more polar than other components will experience _____ attraction to the stationary phase leading to a _____ retention time.

A less, shorter

B greater, longer

Question 7

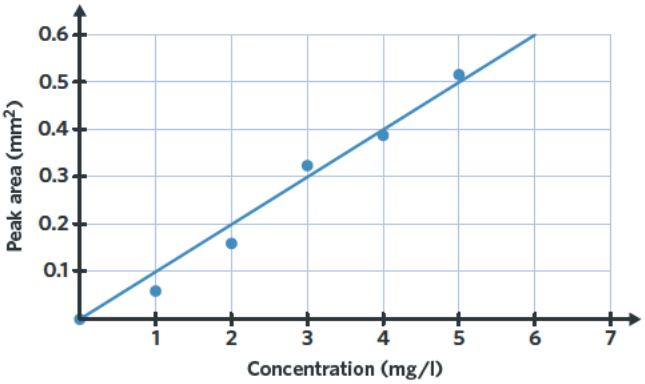
Chromatography can be used to analyse organic compounds.

A True

B False

Question 8

The calibration curve for a component in a chromatography column is shown in the following graph.



If a sample was run and produced a peak area of 0.3, which of the following concentrations would be closest to the true value?

A 2.5 mg/l

B 3 mg/l

Question 9

Terry knows that the concentration of caffeine in his morning flat white should be around 95 mg/L. In order to create a calibration curve to accurately determine the concentration of caffeine in his flat white by chromatography, he should use samples of

A 10, 20 and 30 mg/L

B 70, 90, 110 mg/L

Question 10

Chromatography is a _____ analytical technique.

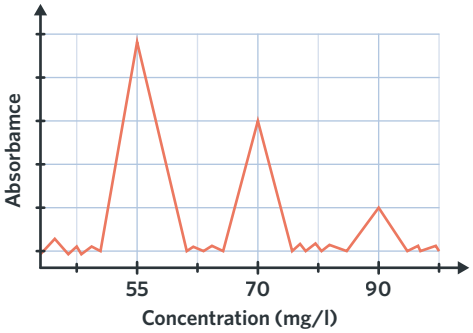
A destructive

B non-destructive

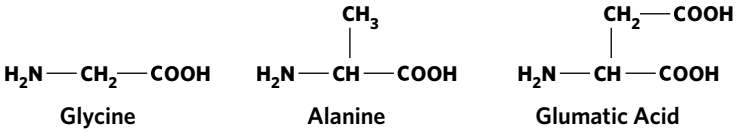
Deconstructed questions

The following statement is to be used for questions 11–13.

Emily is a biochemist and she is investigating the different amino acids present in a solution. She generates the following chromatogram using a non-polar stationary phase and polar mobile-phase. From her studies about amino acids, she knows that amino acids contain polar groups.



She knows that one of three amino acids could be present shown:



Question 11 (1 MARK)

The greatest force of attraction to the stationary phase will be experienced by

A alanine.

B glycine.

C can't distinguish between glycine and alanine.

D glutamic acid.

Question 12 (1 MARK)

The greatest forces of attraction to the mobile phase will be experienced by

A alanine.

B glycine.

C can't distinguish between glycine and alanine.

D glutamic acid.

Question 13 (3 MARKS)

Using the above information, match the three retention time peaks to each of the three amino acids and state which amino acid is present in the largest concentration.

Exam-style questions

Within lesson

Question 14 (1 MARK)

Consider the following statements about a high performance liquid chromatography (HPLC) column that uses a non-polar mobile phase and a polar stationary phase to analyse a solution:

I Polar molecules in the solution will be attracted to the solvent particles by dipole-dipole attraction.

II Non-polar molecules in the solution will be attracted to the stationary phase by dispersion forces.

III Non-polar molecules in the solution will travel through the HPLC column more rapidly than polar molecules.

Which of these statements are true?

A I and II only

B I and III only

C II and III only

D I, II and III

Adapted from VCAA 2015 Exam Section A Q8

Question 15 (1 MARK)

A chemistry student has inadvertently mixed benzene (non-polar compound) and water together. Which of the following HPLC columns would lead to water eluting from the column in a shorter time than benzene?

	Stationary phase	Mobile phase
A	Octane (C_8H_{18})	Ethanol
B	Silica	Propane
C	Silica	Hexane
D	Silica	CCl_4

Question 16 (1 MARK)

An ecologist wants to see the effect that the establishment of a clothing factory next to a river has had on the ecosystem. It is decided that samples from the river will be analysed using HPLC. In order to determine the effect on the ecosystem, both the concentration and identity of components in the sample need to be identified. What is the simplest way for this to be achieved?

A Only run the water sample in the HPLC column

B Run samples of possible contaminants in the HPLC column and then run the water sample in the HPLC column

C Run samples of possible contaminants in the HPLC column, run samples of known concentrations in the HPLC column and then run the water sample in the HPLC column

D Run samples of possible contaminants in the HPLC column, run samples of known concentrations in the HPLC column, generate a calibration curve and then run the water sample in the HPLC column

Question 17 (9 MARKS)

A scientific researcher has just purchased a new HPLC device and wanted to start using it to determine the concentration of a specialised immune cell in the blood of a patient recently diagnosed with COVID-19.

a Using your knowledge of chromatography, explain how components of a sample are separated by HPLC. (3 MARKS)

b Outline the steps required to produce a calibration curve for this experiment. (2 MARKS)

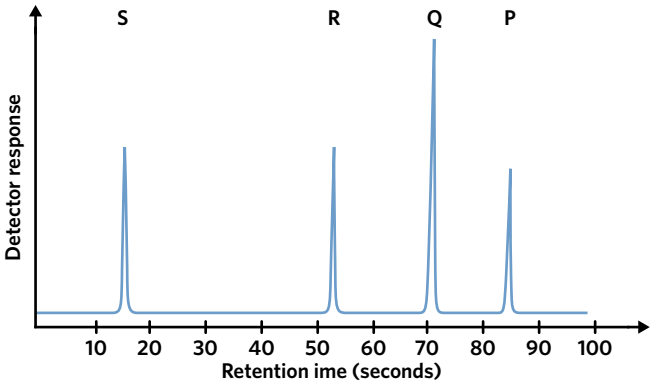
A calibration curve was then generated. However, a few moments later, a laboratory assistant opened a window in the lab, and the temperature of the room increased.

c What will happen to the retention times of components of a sample at this new elevated temperature. (2 MARKS)

d If the concentration of this immune cell needs to be determined accurately, will a new calibration curve have to be generated? Explain why or why not. (2 MARKS)

Question 18 (7 MARKS)

A small sample of blood was taken from a crime scene and sent off to the laboratory for forensic analysis. The coroner decided to use HPLC with a polar stationary phase to analyse the sample. The results are shown.



Adapted from VCAA 2016 Short Answer Question 1

a Based on the chromatogram, how many components are in the sample? (1 MARK)

b Which of component S, P, Q, R is the most polar? Explain your answer (2 MARKS)

c Suggest how the components S, P, Q, R could be accurately identified in order to help the police solve the crime. Explain why this is possible with the analytical technique of chromatography. (2 MARKS)

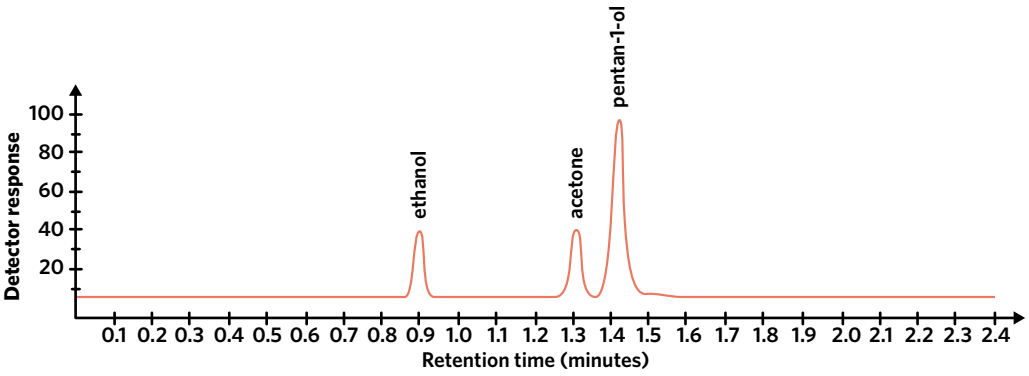
Police soon suspect that alcohol was involved in the crime and instructed the coroner to determine the concentration of alcohol in the sample of blood.

d Explain how the coroner could, given they know which peak corresponds to ethanol, determine the concentration of alcohol in the sample of blood. (2 MARKS)

Multiple lessons

Question 19 (9 MARKS)

An innovative police officer decided to test the potential usage of high performance liquid chromatography columns in booze buses. A sample of a driver's saliva was taken and analysed.



Adapted from VCAA 2010 Chemistry Exam 1 Short Answer Question 5

a Which molecule, pentan-1-ol or ethanol, is more polar? Explain your answer. (2 MARKS)

b Based on the above, has a polar or non-polar stationary phase been used? Justify your answer. (2 MARKS)

c Is this chromatogram sufficient in order to determine whether the driver is below or above the limit of 0.05%? Explain (2 MARKS)

In practice, booze buses do not use high performance liquid chromatography to analyse the breath of drivers.

d Suggest three limitations for the use of HPLC in booze buses. (3 MARKS)

Key science skills

Question 20 (7 MARKS)

Charles, a budding entrepreneur, recently created his own brand of tea, sourced sustainably from Sri Lanka. He claims that each cup of 100 mL contains 80 mg of caffeine. In order for this information to be published, the department of food safety decided to use chromatography to confirm his claims. Samples of known concentrations of caffeine were run through the machine and the area under the peaks recorded. The following table was generated.

Concentration (mg/100mL)	Area under peak (arbitrary units)
20	103
30	148
40	200
50	247
60	295
Charles' tea	180

a What is the dependent and independent variable in the generation of the calibration curve? (2 MARKS)

b Construct a calibration curve based on the data, labelling the axes appropriately. (2 MARKS)

c Comment on the accuracy of Charles' claim. (3 MARKS)

Exam-style question hints

14: All molecules are able to form instantaneous dipoles. **15:** Retention time depends in the polarities of the mobile phase, stationary phase and components in the sample. **16:** Retention times and areas under peaks can only be useful for identification if the samples have been run in the HPLC column under identical conditions. **17a:** Chromatography relies on different strengths of attraction to the stationary and mobile phase. **17b:** A calibration curve links the concentration of a component to its absorbance. **17c:** Temperature affects solubility and kinetic energy. **17d:** Values can only be compared if subjected to identical conditions. **18a:** The peaks in a chromatogram reflect the components in a sample. **18b:** The polarity of the stationary phase, mobile phase and a component, affects the retention time of a component. **18c:** Different compounds have different retention times. **18d:** A calibration curve links the concentration of a component to its absorbance. **19a:** The effect of a polar functional group depends on the size of the molecule. **19b:** The polarity of the stationary phase, mobile phase and a component, affects the retention time of a component **19c:** Chromatograms can only determine concentrations with calibration curves. **19d:** It is important to consider the practical limitations of an analytical technique **20a:** Independent variable is what is being changed whereas the dependent variable is what is being measured **20b:** A calibration curve requires a line of best fit **20c:** Accuracy depends on the closeness of the measured value to the true value.

EXPERIMENT

NEUTRALISATION OF SOFT DRINKS

Soft drinks often contain citric acid ($C_6H_8O_7$), which can have negative health effects such as tooth decay. To determine the acidity of soft drinks, a titration experiment can be performed using a strong base such as sodium hydroxide (NaOH).

Aim: To determine the concentration of oxalic acid in lemonade.

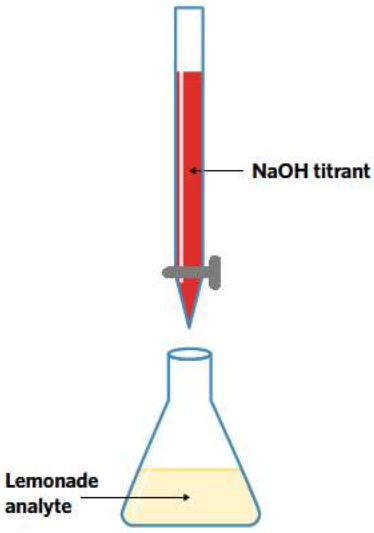
Materials

(per group)

- 1 × 25.0 mL pipette
- 1 × 50.0 mL burette
- 1 × Burette stand and clamp
- 1 × 250 mL conical flask
- 1 × 250 mL beaker
- 1 × Funnel
- 100 mL of lemonade (any brand; opened a few days prior to release CO_2)
- 1 × Phenolphthalein indicator
- 200 mL of 0.10 M sodium hydroxide, NaOH

Method

1. Add 200mL of 0.10 M NaOH into a beaker.
2. Pipette a 25.0 mL aliquot of lemonade into a conical flask.
3. Add three drops of phenolphthalein indicator to the lemonade sample.
4. Use the funnel to fill the burette with 0.10 M NaOH.
5. Record the volume of NaOH in the burette.
6. Slowly and gradually add NaOH into the conical flask until the indicator permanently changes from colourless to pink.
7. Record the volume of NaOH in the burette.
8. Repeat steps 2–7 until three concordant titres are obtained.



Results

	Trial 1	Trial 2	Trial 3	Trial 4
Initial burette reading (mL)				
Final burette reading (mL)				
Titre volume (mL)				

QUESTIONS

Question 1 (1 MARK)

Citric acid is a triprotic acid, capable of donating three hydrogen ions. Write the full balanced equation (including states) for the reaction between citric acid and sodium hydroxide.

Question 2 (5 MARKS)

A student conducted the experiment and recorded the following titres.

a Identify the three titre values that the student should use as concordant titres and justify your answer. (2 MARKS)

Titre 1	Titre 2	Titre 3	Titre 4	Titre 5	Titre 6
30.24 mL	25.52 mL	26.19 mL	30.45 mL	25.43 mL	25.47 mL

b Using your answer from the previous part of this question, calculate the concentration of citric acid in the sample. (3 MARKS)

Question 3 (3 MARKS)

Using the results collected in your own experiment, calculate the concentration of citric acid in your sample.

Question 4 (6 MARKS)

a For each apparatus, identify the correct solution/substance from the above experiment that should be used to rinse the apparatus in between trials. (3 MARKS)

	Solution used to rinse
Burette	
Conical flask	
Pipette	

b Whilst in a rush, a student washed the burette with deionised water. Describe the effect of this on the results of the experiment. (3 MARKS)

ANSWERS

1 $C_6H_8O_{7(aq)} + 3NaOH_{(aq)} \rightarrow C_6H_5O_7Na_{3(aq)} + 3H_2O_{(l)}$

2 **a** [Concordant titres are titres within a range of 0.10 mL.[1]] [Since the titres of volume 25.43 mL (Titre 5), 25.47 mL (Titre 6) and 25.52 mL (Titre 2) are within a range of 0.09mL, these three values would constitute concordant titres.[2]]

✓ ✗ I have defined the term concordant results.[1]

✓ ✗ I have identified the concordant results.[2]

b Average titre $= \frac{25.52 + 25.43 + 25.47}{3} = 25.47$ mL $= 0.02547$ L

$n(NaOH) = 0.10 \times 0.02547 = 0.002547$ mol

$n(NaOH) : n(C_6H_8O_7) = 3 : 1$

$n(C_6H_8O_7) = \frac{1}{3} \times 0.002547 = 0.000849$ mol

$c(C_6H_8O_7) = \frac{0.000849}{\left(\frac{25.00}{1000}\right)} = 0.034$ M

3 Please have your answer checked by your teacher.

4 **a**

Apparatus	Solution used to rinse
Burette	NaOH (titrant)
Conical flask	Water
Pipette	Lemonade (analyte)

b [If the burette is washed with deionised water, the 0.10M NaOH titrant solution would be diluted to an unknown extent.[1]] [Consequently, a greater volume of NaOH would be required to neutralise the citric acid.[2]] [The increased titre volume would mean that the citric acid concentration calculated would be larger than the true value.[3]]

✓ ✗ I have identified the effect of using water in the burette.[1]

✓ ✗ I have described the effect on the volume of titrant used.[2]

✓ ✗ I have described the effect on the concentration of citric acid calculated.[3]

CHAPTER 10 REVIEW

MULTIPLE CHOICE QUESTIONS (10 MARKS)

Question 1 (1 MARK)

The majority of water on Earth is

A used for animal agriculture.
B contaminated with industrial waste.
C freshwater found in oceans.
D too saline for human consumption.

Question 2 (1 MARK)

Bioaccumulation is the process by which

A heavy metals build up in predators higher in the food chain.
B heavy metals have an affinity for red blood cells in warm-blooded organisms.
C industrial and agricultural run-off increases metal concentrations in waterways.
D nitrogen, phosphorus and potassium from fertilisers decrease soil fertility over time.

Question 3 (1 MARK)

Water salinity can be tested by

A placing a lightbulb in a water sample.
B connecting a sample to an electrical circuit and measuring the conductivity.
C adding a solid to a solution and determining the extent of dissolution.
D observing the colour of a water sample.

Question 4 (1 MARK)

Consider the following reaction:

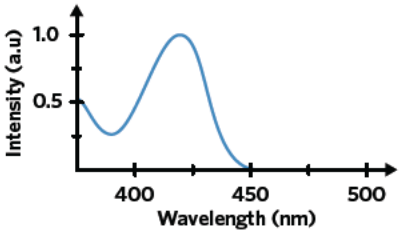
$CaCl_2 + 2KOH \rightarrow Ca(OH)_2 + 2KCl$

If 0.42 g of $CaCl_2$ was allowed to react in excess KOH, calculate the mass of precipitate formed.

A 0.28 g
B 0.0025 g
C 0.38 g
D 0.19 g

Question 5 (1 MARK)

The following UV-Vis spectra was formed from the analysis of a solution.



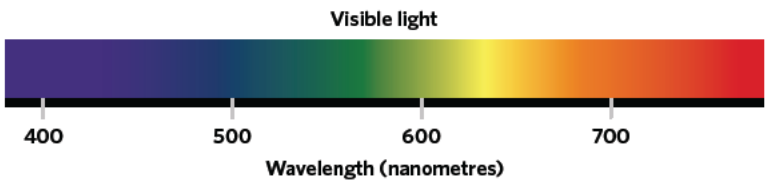


Image: brgfx/Shutterstock.com

If this sample was to undergo colorimetry analysis, the filter used should select for

A a purple colour.
B a yellow/orange colour.
C a red colour.
D a green colour.

Question 6 (1 MARK)

An intensely orange coloured solution was analysed using colorimetry. Which of the following statements is not true?

A The concentration of ions in the solution is inversely proportional to the amount of light detected by the detector.

B If the colour of the solution is as a result of the ion, a darker solution means that more light will be absorbed.

C A blue/violet light is used.

D The solution becomes a lighter orange during analysis.

Question 7 (1 MARK)

A copper ion solution was to be analysed by AAS. Which of the following statements is true for this experiment?

A A copper cathode lamp should be used.

B Light emitted from electrons in copper returning to the ground state is absorbed by a detector.

C A copper ion solution with greater volumes will absorb more light.

D A copper ion solution at higher pressures will absorb greater amounts of light.

Question 8 (1 MARK)

In volumetric analysis, a _____ contains the _____, of precisely known concentration, and a _____ is used to deliver the _____, of unknown concentration.

A pipette, titre, burette, aliquot

B burette, titre, pipette, aliquot

C pipette, aliquot, burette, titre

D burette, aliquot, pipette, titre

Question 9 (1 MARK)

Which of the following is not a necessary criterion for a standard solution in volumetric analysis?

A The substance must be pure and in a stable form.

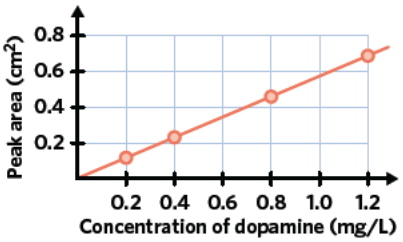
B The substance must be non-reactive with the atmosphere.

C The substance must be inexpensive.

D The substance must have a low molecular weight.

Question 10 (1 MARK)

Using the calibration curve shown, identify the correct concentration (mg/L) of a solution with a peak area of 0.6 cm^2 on a chromatogram.



A 1.05

B 0.80

C 1.25

D 0.70

SHORT ANSWER QUESTIONS (30 MARKS)

Question 11 (5 MARKS)

Mahesh is a chemistry student learning how to conduct volumetric analyses. He wants to determine the concentration of oxalic acid in a sample of carrot juice. The concentration of oxalic acid ($C_2H_2O_4$) in the sample is important as high concentrations may induce vomiting.

Before attempting the analytical technique, he tries to familiarise himself with some key terms.

a Complete the following table to assist Mahesh. (2 MARKS)

Term	Definition
Equivalence point	
End point	
	Titres within a range of 0.10 mL.

b Mahesh adds beetroot juice with an unknown concentration of oxalic acid to a burette and adds 25 mL 1.0 M sodium hydroxide (NaOH) to a pipette.

i Oxalic acid is a diprotic acid, able to donate both of its hydrogen ions. Write the equation of the reaction between oxalic acid and sodium hydroxide. (1 MARK)

ii After carrying out the reaction a number of times, the average titre of beetroot juice was found to be 27.4 mL. Calculate the concentration of oxalic acid in the beetroot juice. (2 MARKS)

Question 12 (9 MARKS)

Antony is a statesman in Rome who has been designated the task of assessing the quality of the city's water supplies. Antony decides to use gravimetric analysis and electrical conductivity testing as potential methods of determining water salinity.

a When carrying out gravimetric analysis, Antony must ensure that he knows which reactant is limiting and which is in excess.

i Should the ion under analysis or the added reactant be the limiting reactant? Justify your answer. (2 MARKS)

ii Identify the limiting reactant in the following reaction:

$Pb(NO_3)_{2(aq)} + CaCl_{2(aq)} \rightarrow PbCl_{2(s)} + Ca(NO_3)_{2(aq)}$ with 6.34 g of $Pb(NO_3)_{2(aq)}$ and 2.51 g of $CaCl_{2(aq)}$.

Show your calculations. (3 MARKS)

iii For the reaction shown in question 12aii, calculate the mass in grams of $PbCl_2$ that will be formed. (2 MARKS)

b Antony also tries to learn about the method of electrical conductivity testing. When comparing a sample of pure water with a river water sample, he notices that the electrical conductivity in the river sample is far greater. Explain this observation. (2 MARKS)

Question 13 (8 MARKS)

Ziggy is a martian investigating the properties of materials on Earth. As the principles of science are consistent throughout all of the observable universe, Ziggy's martian home also makes use of chemical analysis techniques found on Earth, such as colorimetry and atomic absorption spectroscopy. Ziggy scoops up a sample of water from the Yenisei River in Siberia, and compares it to a sample of water taken from the Yangtze River in China.

a Identify two potential sources of contamination in each of these rivers. (2 MARKS)

b Describe the processes of colorimetry and atomic absorption spectroscopy, highlighting the similarities and differences. (3 MARKS)

c Ziggy decides to assess the rivers for the concentration of potassium in each. He possesses a series of standard solutions of fluoride, and records the following:

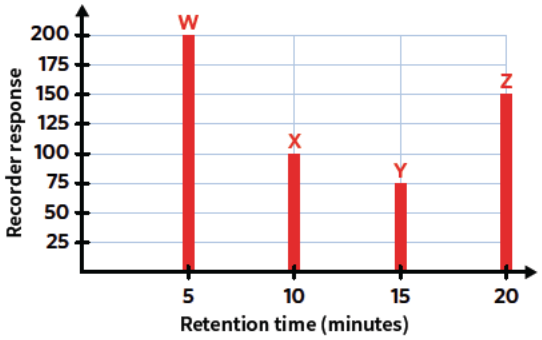
Fluoride concentration (ppm)	0.5	1.0	1.5	2.0	2.5
Absorbance	0.13	0.26	0.40	0.53	0.67

i Plot the data on the graph, and draw a line of best fit. (2 MARKS)

ii In order for water to be safe to drink, the concentration of fluoride must be less than 1.5 ppm. In atomic absorption spectroscopy, the sample from the Yenisei River obtained an absorbance value of 0.37, while the sample from the Yangtze River obtained an absorbance value of 0.64. Explain whether these samples are safe to drink. (1 MARK)

Question 14 (8 MARKS)

Chromatography is an analytical technique that enables scientists to obtain qualitative data about the components of a sample. In some types of chromatography, quantitative data can also be obtained. A researcher uses HPLC with a non-polar stationary phase and polar mobile phase to produce the chromatogram shown.



a According to the chromatogram, how many substances are in the sample? (1 MARK)

b Define the term retention time. (1 MARK)

c Explain which of the components is the most polar. (2 MARKS)

d Identify which of the components has the highest concentration. (1 MARK)

e Fill in the following table to identify impacts of changing retention time of changing conditions. (3 MARKS)

Change to conditions	Effect on retention time
Longer column	
Increased flow rate of mobile phase	
Decreased temperature	

KEY SCIENCE SKILLS QUESTIONS (5 MARKS)

Question 15 (5 MARKS)

Mark is carrying out volumetric analysis. He records the following results.

Titration number	1	2	3
Final burette reading	23.70	28.60	22.20
Initial burette reading	2.50	7.35	1.60

a Is the data collected qualitative or quantitative data? (1 MARK)

b Are these titres concordant? Justify your answer with calculations. (2 MARKS)

c Identify one systematic error and one random error that may produce non-concordant titres. (2 MARKS)

1A Scientific research

Theory review questions

1 B. Scientific ideas are rigorously tested through the scientific method.

2 B. Scientific ideas need to be able to be replicated by others.

3 B. Scientific ideas are based on objective evidence.

4 B. The basic scientific method begins with a brief introduction → method → results → discussion → conclusion.

5 A. Scientific knowledge is built over time on top of previous discoveries.

6 B. Not all conclusions are accurate and therefore need to be reviewed by others.

7 B. Textbooks present information that has been interpreted from original research articles.

8 B. Results are not always in line with theory.

9 B. A hypothesis highlights the potential effect of the independent variable on the dependent variable.

10 B. The independent variable is the only variable that is deliberately changed in an experiment. As a result, this variable causes the changes observed in an experiment.

Deconstructed questions

11 B **12** D

13 The breakdown of the egg shell.

Exam-style questions

Within lesson

14 [The ideas around the Zodiac signs have not been established using the scientific method.[1]][Therefore, the ideas are currently considered to be non-scientific.[2]]

✓ ✗ I have identified the need for the scientific method to classify an idea as scientific or not.[1]

✓ ✗ I have linked my answer to the question.[2]

15 **a** The wavelengths/colours used in the experiment.

b [If the attraction of bugs to light depends on the wavelength of light used,[1]][then the colour the that attracts the most bugs will reflect the wavelength that bugs are most attracted to.[2]]

✓ ✗ I have identified the effect of the independent variable on the dependent variable.[1]

✓ ✗ I have identified a possible outcome based on the relationship between the independent and dependent variable.[2]

1B Conducting an experiment

Theory review questions

1 A. Qualitative analysis gives observational data.

2 B. Quantitative data refers to numerical data.

3 A. Primary data is collected directly by the experimenter.

4 B. Secondary data refers to a collection method, not the type of data that can be collected.

5 B. Findings from research studies can have negative impacts on different areas in society.

6 B. Scientists have to consider all of the possible uses of their research, however are not responsible for how it is used by others.

7 B. Previous work that contributes to research should always be referenced.

8 B. Research findings are generally tested multiple times and shared within the scientific community first so that they can be checked by other experts before it is shared to the wider community.

9 B. Risks associated with substances are already documented in the MSDS.

10 B. Researchers need to take into account potential safety risks prior to conducting research.

Deconstructed questions

11 D **12** B

13 [Although CO_2 increases can be seen in the increase in the rate of crop growth, the idea that an increased concentration of CO_2 can result in a decrease in the amount of nutrients present in the crop can have a negative environmental impact.[1]][To get the same amount of nutrients, animals and humans would need to eat more crops.[2]][This would require more land to be cleared so that it can be used to grow more crops, causing a loss of habitat for many animals.[3]]

✓ ✗ I have identified the key idea from the passage.[1]

✓ ✗ I have identified the effect of the idea.[2]

✓ ✗ I have described the impact of this idea.[3]

Exam-style questions

Within lesson

14 **a** Secondary data

b Plagiarism

15 **a** Qualitative

b [In his experiment, Luis is collecting data based on the different colour changes.[1]][This is classified as qualitative data.[2]][In order to calculate the amount of a substance present, Luis would need to collect quantitative data that gives numerical amounts.[3]][Unless Luis uses an analysis technique that allows the colours and their intensities to be translated into numerical values, Luis won't be able to calculate the amount of substance present by only using the data he collected.[4]]

✓ ✗ I have identified the data collected from the experiment.[1]

✓ ✗ I have identified the type of data collected from the experiment.[2]

✓ ✗ I have described the type of data that needs to be collected to calculate amounts.[3]

✓ ✗ I have identified what is required for Luis to calculate the amount of substance present.[4]

16 **a** Ethical

b [As vaccines are generally produced for human use, testing on humans would allow for the development of an effective vaccine in a shorter period of time.[1]][As a result, the vaccine would be available to affected individuals in a shorter period of time.[2]]

I have identified the purpose of vaccines/medicine.[1]

I have described the impact of human testing.[2]

c [Although human experimentation allows us to develop vaccines quickly using less resources, there is a chance that the individuals that are involved in the testing can become sick themselves.[1]] [The alternative testing method that does not involve humans takes a significant amount of time, which could not only have high economic costs, but also means that those affected by disease may not be able to receive the appropriate treatment in time.[2]] [When choosing a method, it is important that we consider and understand all of the immediate and secondary implications of the method.[3]]

I have identified a disadvantage of human testing.[1]

I have described the disadvantages of alternative methods.[2]

I have identified the considerations when choosing an appropriate method.[3]

17 a Lab coat & safety goggles

b [Different experiments use different reagents, which carry their own risks.[1]] [As a result, experimenters are required to develop a risk assessment specific to their experiment.[2]] [As a result, it is the experimenter and not the lab technician that should be writing a risk assessment.[3]]

I have identified the nature of experimental risk.[1]

I have identified the role of the experimenter in the risk management of an experiment.[2]

I have linked my answer to the question.[3]

Multiple lessons

18 a [If the level of sugar in drinks increases then the tax on manufacturers will increase.[1]] [To avoid costs, there will be a decrease in the proportion of drinks with over 5 g sugar per 100 mL (which incurs no charge).[2]]

I have referred to the independent variable in the hypothesis.[1]

I have described the effect of the independent variable on the dependent variable.[2]

b External ethical concern

c [As mentioned in the passage, sugar can have negative health implications such as obesity and diabetes.[1]] [The SDIL can create an incentive for the company to reduce the amount of sugar in their drinks.[2]] [This means that there would be more drink options with less sugar available for people to consume.[3]] [Therefore, the number of people with sugar-related disease would decrease.[4]]

I have identified an impact.[1]

I have described the effect of SDIL.[2]

I have explained the effect of the SDIL on the availability of drinks with less sugar.[3]

I have identified the overall impact on population health.[4]

Key science skills

19 a The brands of bottled water

b
- The volume of water used
- The temperature of the water
- The type of technique used for analysis

Other answers may include:

- The types of reagents used
- The amounts of reagent used

1C Interpreting data

Theory review questions

1 B. Experimental data is represented in a way that showcases the data best.

2 A. Qualitative and quantitative data are better represented in different ways.

3 A. Without values, pie charts show relative comparisons.

4 B. The point of scientific research is to identify the effect of the independent variable on the dependent variable.

5 A. For the graph, the x-axis represents the independent variable.

6 B. The dependent variable is affected by the independent variable.

7 B. Control groups are generally not subjected to the independent variable.

8 B. Not all relationships seen in experiments are causal.

9 B. Other variables can impact the experiment, and therefore the conclusion drawn.

10 B. The supporting or refuting of a hypothesis depends on the results obtained in the experiment.

Deconstructed questions

11 C **12** A

13 [At the beginning of the experiment, the pH of the solution is close to 0.[1]] [Therefore the original solution can be considered acidic, as acids have a low pH.[2]]

I have used the graph to determine the pH of the original solution.[1]

I have explained why the solution would be considered as acidic.[2]

Exam-style questions

Within lesson

14 a As the reaction proceeds, the mass of the beaker and its contents remain the same.

b [According to the law of conservation of mass, no mass is lost or gained in a chemical reaction.[1]] [Therefore, for the reaction in the experiment which is sealed, there would be no loss in mass recorded as the reaction proceeds.[2]] [This can be shown by the data presented in the graph, where there was no change in mass over time.[3]]

I have identified the underlying chemical concept.[1]

I have described the effect of the chemical concept.[2]

I have used data to justify my response.[3]

15 **a** According to the graph, the world derives most of its electrical energy from fossil fuels, followed by hydro, nuclear, wind, other renewable resources and solar.

b [The use of fossil fuels both disrupts the earth (due to drilling) and produces greenhouse gases such as carbon dioxide, as well as dangerous compounds such as sulfur dioxide and nitrogen oxides.[1]] [Due to the growing population, it is expected that we will need to use more fuels to provide more energy to more people.[2]] [This could result in negative environmental impacts.[3]] [As such, we will need to be shifting to more environmentally friendly fuel alternatives.[4]]

- ✓ ✗ I have identified the main factors involved in the use of fossil fuels.[1]
- ✓ ✗ I have described the impact of a growing population on energy consumption.[2]
- ✓ ✗ I have identified a major impact of the use of energy resources over time.[3]
- ✓ ✗ I have identified an alternative to the use of fossil fuels.[4]

Multiple lessons

16 **a** Quantitative data

b Different energy drinks contain different levels of caffeine.

c [There are major health concerns related to the consumption of energy drinks.[1]] [Hence it is inappropriate to have these drinks readily available for and tested on teenagers.[2]]

- ✓ ✗ I have identified the negative effects related to energy drinks.[1]
- ✓ ✗ I have identified an ethical consideration.[2]

d Control group consumes non-caffeinated energy drinks.

17 **a** Number of colours (in the texta)

b Answers may include:

- Amount of texta used
- Brand of texta
- The amount of time the experiment was conducted for
- The type of paper used (to test)

1D Experimental factors affecting data

Theory review questions

1 A. Spilling a solution is a random error.

2 A. Forgetting a reagent is a random error.

3 B. Random errors do not affect results in consistent ways.

4 A. Precise data are within close proximity of each other.

5 A. Values that are accurate are close to the theoretical value.

6 A. The zeros after the first significant figure are considered significant.

7 A. Addition requires answers to be expressed to the lowest number of decimal points in the calculation.

8 B. Multiplication requires answers to be expressed to the least number of significant figures in the calculation.

9 A. Reproducibility describes results collected by different experimenters.

10 A. In an experiment, only the independent variable should be changed.

Deconstructed questions

11 D **12** D

13 [For the experiment to be valid, all variables outside the independent variable should be controlled.[1]] [For this experiment, the independent variable is the conditions in which the banana was kept.[2]] [Despite this, many other variables were also changed in the experiment. This included the size of the bananas and the time intervals between measurements.[3]] [Therefore the conclusion isn't very valid as other variables may have affected the results obtained.[4]]

- ✓ ✗ I have identified the parameters of validity.[1]
- ✓ ✗ I have identified the independent variable.[2]
- ✓ ✗ I have identified the changes in other variables.[3]
- ✓ ✗ I have linked my answer to the question.[4]

Exam-style questions

Within lesson

14

Description	Type of error
Experimenter uses an uncalibrated thermometer to measure the heat of a reaction.	Systematic
Some reagent falls off a measuring glass as a student walks back to the bench.	Random
Student reads a measuring cylinder from the top of the meniscus.	Parallax
Student uses the wrong reagent in a single reaction.	Personal

15 **a** Tests 1, 2 and 4.

b [It was expected that the results would be around 15.00 mL of HCl.[1]] [However, most of the results collected were much higher than this value.[2]] [Although there was one result that was accurate, the average of all the results, 17.8 mL, is still higher than the true value.[3]] [Therefore, the results are not very accurate.[4]]

- ✓ ✗ I have identified the expected results.[1]
- ✓ ✗ I have identified the main difference between the expected and actual results.[2]
- ✓ ✗ I have described the the data collected.[3]
- ✓ ✗ I have linked my answer to the question.[4]

c Two

16 **a**
- Accessibility of reagents
- The quantity available for testing

b [Reproducibility is a measure of how close results are of the same experiment being conducted in a different environment by different experimenters.[1]] [Due to the fact the results obtained in the experiments were different,[2]] [the experiment is not considered to be very reproducible.[3]]

- ✓ ✗ I have defined reproducibility.[1]
- ✓ ✗ I have identified the key difference between the experiments.[2]
- ✓ ✗ I have linked my answer to the question.[3]

c Use more precise equipment.

d [Reproducibility describes the ability of results to be replicated using the same method but conducted in different conditions by different experimenters.[1]][Repeatability, however, measures the ability of results to be replicated when conducted under the same conditions by the same experimenter.[2]]

✓ ✗ I have defined reproducibility.[1]

✓ ✗ I have compared reproducibility to repeatability.[2]

Multiple lessons

17 **a** % of fat in the chocolate.

b Chocolate sample with no fat.

c As the % of fat in the chocolate increases, the melting temperature decreases.

d Three

18 **a** The change in blood glucose levels.

b The higher the sugar content in food, the higher the blood glucose level.

c [Thomas's glucose meter was consistently measuring the glucose levels 0.20 mmol/L less than the true value.[1]][This would affect the accuracy of his results and is a systematic error.[2]]

✓ ✗ I have identified the key factor affecting the results.[1]

✓ ✗ I have linked my answer to the question.[2]

1E Writing scientific material

Theory review questions

1 B. Clear writing allows for the findings to be easily understood.

2 A. It is better to be concise in scientific writing.

3 B. A methods section is required in all scientific papers.

4 B. Chemical terminology can represent a large concept that contains many sub-ideas.

5 B. Units depend on the information being expressed.

6 B. Measuring distance requires a unit that can indicate a distance.

7 A. Chemical equations represent a process involved in chemical reactions.

8 B. Visual models represent a concept/idea.

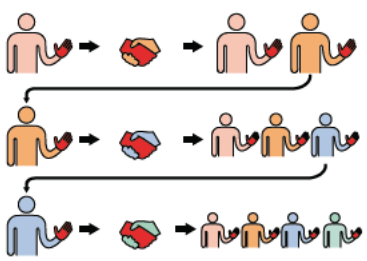
9 B. Models cannot show all of the details of a concept.

10 B. The details regarding how each heat transfer method operates aren't shown in the model.

Deconstructed questions

11 B **12** C

13



✓ ✗ I have shown the process of the spread of germs through people.

Exam-style questions

Within lesson

14 **a** Different colours of light have different wavelengths/patterns.

b nm

c We are unable to see why each colour has different waves.

Multiple lessons

15 **a** Days

b Test A

c The length of carrot

16 **a** Size of rock used

b To test the effect of the use of different sized rocks on the size of crystals that are able to grow.

c Title, introduction, materials and methods, results, discussion, conclusion and references.

Chapter 1 review

Multiple choice questions

1 A. Qualitative data is based on descriptive information.

2 B. In this experiment, the relationship between the independent and dependent variables is correlated (as long as the experiment is appropriately controlled).

3 C. Controls show that any changes in the dependent variable were due to the independent variable and cannot improve the accuracy of the results.

4 B. Errors in reading measuring equipment affect all readings in the same way, however results can still be measured as being close to each other.

5 B. There are factors that can limit the ability of an experiment to be conducted. Errors are not considered as limitations.

6 D. MSDS provides information regarding the hazards of substances and techniques used.

7 C. Random errors are unpredictable and occur irregularly.

8 C. All variables except the independent and dependent variable need to be controlled.

9 D. Data represented in a graph shows a pattern between the independent and dependent variable.

10 B. Precise data are within close proximity of each other.

Short answer questions

11 **a** Amount of corn flour.

b Qualitative data.

c [As shown in the table, every sample contains a different amount of corn flour and water.[1]][The experimenter has changed multiple variables (corn flour and water), and therefore the results are not valid[2]][as the observations would not be able to be attributed to the independent variable (corn flour) alone.[3]]

✓ ✗ I have identified the change in multiple variables.[1]

✓ ✗ I have identified the relationship between changing variables and validity.[2]

✓ ✗ I have described the impact of changing multiple variables.[3]

12 a It is expected that the extra tax placed on unhealthy foods would encourage people to consume healthier food(s).

b [The taxes places on unhealthy foods would mean that the foods would cost more to purchase.[1]][As a result, people would be less likely to buy these foods, leading to a decrease in the consumption of unhealthy foods.[2]]

✓ ✗ I have identified the impact of taxes on food prices.[1]

✓ ✗ I have described the effect of taxes on the consumption of unhealthy foods.[2]

c [By increasing food prices, this could make it more difficult for people to afford and therefore access foods.[1]][This could have the opposite effect, where people could become less healthy due to not being able to eat enough food.[2]][Also, introducing this tax is effectively trying to influence human behaviour, whereby the government is making decisions on what people can and cannot access. This could be considered highly unethical.[3]]

✓ ✗ I have identified the impact of taxes on food accessibility.[1]

✓ ✗ I have identified an adverse impact on individual health.[2]

✓ ✗ I have described an ethical concern related to the Government influencing human behaviour.[3]

13 a A bubble solution that does not contain any additional substances.

b i Line graph

ii Corn syrup

iii [Reproducibility measures the ability of an experiment to collect similar results when conducted by different experimenters.[1]][In this experiment, there is a lot of variability in the way that the data is collected. For example, when taking the photo that is used to measure bubble size, the positioning of the experimenter relative to the bubbles could impact the size of the bubbles measured.[2]][Considered the results recorded for both glycerin and corn syrup are relatively close to each other, this could affect the overall conclusion of the experiment.[3]][Also, the concentration of the solutions and the size of the wand were not included, therefore the experiment cannot be replicated properly.[4]][As a result, the results are not very reproducible.[5]]

✓ ✗ I have defined reproducibility.[1]

✓ ✗ I have described an area of variability in the experiment.[2]

✓ ✗ I have identified an impact of the variability on the conclusion drawn.[3]

✓ ✗ I have described the impact of the exclusion of key information related the materials used.[4]

✓ ✗ I have linked my answer to the question.[5]

iv The use of corn syrup produces the largest bubbles, followed by glycerin and dish washing detergent.

14 a [The law of conservation of energy states that energy is transferred or transformed, which does not mean that the energy needs to be used for a specific purpose.[1]][The energy available in the coal was still transformed into thermal energy, even if some escaped to heat up the surrounding air rather than the food.[2]][If the total energy before and after was measured, even in its different forms, it would be the same.[3]][Therefore, the student's statement is inaccurate.[4]]

✓ ✗ I have identified the main error in the statement relative to the law of conservation of energy.[1]

✓ ✗ I have described the energy change(s) during this scenario.[2]

✓ ✗ I have identified a strategy that can be used to confirm the law of conservation of energy.[3]

✓ ✗ I have linked my answer to the question.[4]

b i Answers may include:

- Being able to measure the energy lost to the environment
- Being able to measure the energy transferred into the food
- Being able to measure the chemical energy in the coal
- The availability of coal / food for the experiment

ii Answers may include:

- The type of coal used
- The type of food used
- The barbecue used
- The distance between the food and the coal
- The length of time of cooking/burning coal

15 a The concentration of vinegar used.

b i [This test is being conducted on humans,[1]][where the effect of the experiment may cause the subjects some discomfort.[2]]

✓ ✗ I have identified a key ethical consideration.[1]

✓ ✗ I have identified the effect of conducting research in this way.[2]

ii [Since the results were derived from many different individuals, the self rating used is subjective and would be based on the previous experiences of each subject.[1]][As a result, the interpretation of the results would vary between subjects.[2]][Also, facial expressions from each subject in response to eating lemon would not be an accurate indicator of sourness as this would also vary between subjects and their experiences.[3]][As a result, the results from this experiment are are not valid.[4]]

✓ ✗ I have identified the subjective of the self rating system.[1]

✓ ✗ I have identified the impact of the self rating system.[2]

✓ ✗ I have described the subjective nature of the use of facial expressions.[3]

✓ ✗ I have linked my answer to the question.[4]

iii There's no direct relationship between the consumption of vinegar and a reduction in a person's reaction when consuming lemons.

2A The atom

Theory review questions

1 A. A scale measures the relative size of particles rather than the actual size.

2 B. The nanoscale relates to very small objects and therefore, as per the data book, the multiplying prefix is 10^{-9}.

3 B. The model of the atom is constantly revised with new discoveries.

4 A. Protons are positively charged, electrons are negatively charged, neutrons are neutrally charged, nucleons are protons and neutrons.

5 A. A lattice is a three-dimensional arrangement of atoms.

6 A. Nanoparticles can function on cancer cells due to their size.

7 A. More than one electron can occupy an energy shell.

8 A. The Bohr model was replaced not due to observing orbitals (which cannot be observed) but rather due to new high resolution spectra showing inconsistencies.

9 A. Orbitals are created by solving a mathematical equation and do not exist in reality.

10 B. Copper and chromium are exceptions to Schrödinger's rule because they want to have every d orbital either partially filled, or fully filled, for stability.

Deconstructed questions

11 A **12** D

13 [The order of energy levels for electron shells is 1s<2s<2p<3s<3p and so on.[1]][Since there are 15 electrons and each s subshell can take two electrons and each p subshell can take six electrons, the electron configuration will be $1s^2 2s^2 2p^6 3s^2 3p^3$.[2]]

✓ ✗ I have identified the order of electron shell energy levels.[1]

✓ ✗ I have given the electron configuration of sulphur according to Schrödinger's model.[2]

Exam-style questions

Within lesson

14 A **15** B **16** B

Multiple lessons

17 **a** **i** $1s^2 2s^1$ **ii** $1s^2 2s^2 2p^6$

iii $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ **iv** $1s^2 2s^2 2p^6 3s^1$

v $1s^2 2s^2 2p^6 3s^2 3p^5$

b **i** Lithium: 2,1 **ii** Neon: 2,8

iii Calcium: 2,8,8,2 **iv** Potassium: 2,8,1

v Chlorine: 2,8,7

18 **a** The excited state

b [A line is produced on the emission spectrum when an electron emits energy after going from the excited state to the ground state.[1]][Since there are four lines on the emission spectrum this means that the electron has transitioned from a higher energy level shell four times.[2]][If the electron starts at shell $n = 1$, this means that the maximum shell number it reaches is $n = 5$.[3]]

✓ ✗ I have identified how a line is produced on the emission spectrum.[1]

✓ ✗ I have linked the number of lines to the number of electron transitions.[2]

✓ ✗ I have identified the maximum shell number that it reaches.[3]

c Three limitations are:

- Bohr's model of the atom is only accurate for single-electron atoms.
- Bohr's model of the atom doesn't take into account the wave-particle duality of small objects.
- Bohr's model doesn't explain the presence of some blurred lines on the emission spectrum.

d [The most recent model of the atom is Schrödinger's model.[1]][Schrödinger's model is different to Bohr's model because it describes the probability of electrons being in certain regions (orbitals) rather than in shells with fixed energy levels as in the Bohr model.[2]][Schrödinger's model also includes subshells which are not described in the Bohr model.[3]]

✓ ✗ I have identified the newest model of the atom.[1]

✓ ✗ I have explained the difference between orbitals and shells.[2]

✓ ✗ I have described the inclusion of subshells in Schrödinger's model of the atom.[3]

19 **a** [Quantum mechanics describes objects which are on the quantum scale (i.e. very small).[1]][This is suitable for describing electrons due to their small size and light mass.[2]]

✓ ✗ I have described the application of quantum mechanics.[1]

✓ ✗ I have explained why quantum mechanics is adequate for describing electrons.[2]

b The orbitals differ in their size, shape, and maximum number of electrons that can be held.

c [The student is incorrect because the 4s orbital is lower in energy than the 3d orbital.[1]][Therefore, according to the Aufbau principle, the 4s orbital will fill before the 3d orbital.[2]]

✓ ✗ I have identified the energy levels of each orbital.[1]

✓ ✗ I have referenced the Aufbau principle to explain the filling of electron shells.[2]

d [Copper: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
Zinc: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$[1]][Even though these atoms have a different number of electrons, they have the same number of electrons in the 3d-subshell as copper is an exception to the Aufbau principle.[2]]

✓ ✗ I have written the electron configurations of copper and zinc.[1]

✓ ✗ I have explained the similarity between copper and zinc.[2]

Key science skills

20 **a** The independent variable is whether the sunscreen has nanoparticles or not.

b Formula for sunscreen (excluding the presence of nanoparticles), amount of sunscreen, intensity of sunlight the beads are exposed to, time spent in the sun, how much time each sample is exposed to the sun for.

c Nanoparticles can absorb a lot of radiation due to their large surface-area to volume ratio.

d Nanoparticles can enter the circulatory system and affect the functioning of the body due to their size.

2B Defining elements

Theory review questions

1 B. The atomic number is the number of protons in the nucleus of an atom.

2 B. The mass number is calculated by the number of nucleons in an atom.

3 B. For an atom to be neutral the positive charge of protons needs to cancel out the negative charge of electrons.

4 B. Isotopes of the same element have the same number of protons, but a different number of neutrons.

5 A. Elements on the periodic table are ordered horizontally (from left to right) in order of increasing atomic number.

6 A. The electron has a significantly lower mass than both the proton and neutron.

7 A. The mass number is equal to the sum of protons and neutrons in any individual atom.

8 A. Mass number does not include the number of electrons.

9 B. Isotopes vary in the number of neutrons.

10 B. Some isotopes can be radioactive.

Deconstructed questions

11 B **12** C

13 [The mass number is the sum of protons and neutrons in an atom.[1]] [Based on the diagram, the image on the left contains a total of nine protons and 10 neutrons, therefore would have a mass of 19. The diagram on the right shows an atom with 10 protons and 10 neutrons, therefore would have a mass number of 20.[2]] [Although both contain the same number of neutrons, they both have a different mass number due to the number of protons in the nucleus.[3]]

✓ ✗ I have defined mass number.[1]

✓ ✗ I have determined the mass number of both atoms.[2]

✓ ✗ I have described the key difference in mass number of the atoms.[3]

Exam-style questions

Within lesson

14 Different number of neutrons in the nucleus.

15 [The atom contains six protons and six neutrons in its nucleus.[1]] [If it were to undergo neutron release, the atom would lose a neutron to become an isotope.[2]] [The atom would then only contain five neutrons instead of six, compared to the original atom.[3]]

✓ ✗ I have identified the particles in the nucleus.[1]

✓ ✗ I have described the effect of neutron release.[2]

✓ ✗ I have compared the nucleus of the atom before and after neutron release.[3]

16 **a**

Elemental symbol	Mass number
R	34
I	36
T	37
A	39

b I and T

c [Considering elements I and T are isotopes of each other, they have the same atomic number.[1]] [As such, they should have the same elemental symbol, and therefore the table should be changed so that both elements share the same elemental symbol.[2]]

✓ ✗ I have identified the key characteristic of isotopes.[1]

✓ ✗ I have described the change that needs to be made to the table.[2]

d $^{39}_{19}A$

e [Both molecules and compounds require the bonding of two atoms.[1]] [If an atom of the element R and A were to bond together, this would mean that there are two atoms involved in RA. As such, it could be defined as a molecule.[2]] [However, considering R and A are two different elements and there may be more than one of each type of atom, RA can also be defined as a compound.[3]]

✓ ✗ I have identified the common characteristic of both molecules and compounds.[1]

✓ ✗ I have explained how RA can be considered a molecule.[2]

✓ ✗ I have described how RA can be considered a compound.[3]

Multiple lessons

17 **a** [The atomic number of an atom indicates the number of protons in an atom, therefore this atom would contain 35 protons.[1]] [Given that the mass number is 81, then the number of neutrons in the atom would be 81 − 35 = 46 neutrons.[2]] [Considering the atom is neutral, the number of electrons would be equal to the number of protons. As such, there would be 35 electrons in that atom.[3]]

✓ ✗ I have described how to the number of protons can be determined.[1]

✓ ✗ I have described how to the number of neutrons can be determined.[2]

✓ ✗ I have described how to the number of electrons can be determined.[3]

b $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

18 **a** 3

b 11

c 23

Key science skills

19 **a** By calibrating the scales, Matthew is minimizing the potential for systematic errors.

b [The mass number of an element is the total number of protons and neutrons in the atom.[1]][Since every element has a different atomic number, it would have a different number of protons compared to other elements.[2]][However, elements can also exist as isotopes, which means that the mass number of elements can still be the same given that the number of neutrons can be different.[3]] [As a result, Matthew's statement is inaccurate.[4]]

- ✓ ✗ I have defined mass number.[1]
- ✓ ✗ I have described the difference in protons between elements.[2]
- ✓ ✗ I have described the impact of isotopes on mass number.[3]
- ✓ ✗ I have linked my answer to the question.[4]

2C The periodic table - part 1

Theory review questions

1 B. Elements in the same period contain the same number of electron shells.

2 A. Elements in the same group have the same number of valence electrons.

3 A. Groups and periods are seen as columns and rows respectively on the periodic table.

4 B. Elements across a period have an increasing number of protons with the same number of inner shell electrons.

5 B. The core charge is the same for all elements in the same group regardless of the atomic size.

6 B. Atomic radius decreases across a period.

7 B. An increase in core charge causes elements in the same row to have a smaller atomic radius.

8 B. Elements towards the top of a group have a smaller atomic radius, and therefore the valence electrons are held closer to the nucleus.

9 A. The closer the valence electrons are to the nucleus, the more difficult it is to remove a valence electron.

10 B. Electronegativity describes the ability of an atom to attract electrons to itself.

Deconstructed questions

11 D **12** D

13 [Both elements have six valence electrons, however sulfur only has three electrons shells compared to four in selenium.[1]][As a result, the valence electrons in selenium are further from the nucleus than sulfur and therefore the attraction between the nucleus and valence electrons becomes weaker.[2]][Therefore, as we move down a group, less energy is required to remove valence electrons,[3]][causing selenium to have a lower ionisation energy than sulfur.[4]]

- ✓ ✗ I have identified the similarity between both elements.[1]
- ✓ ✗ I have compared the electron arrangement of both elements.[2]
- ✓ ✗ I have described the effect of electron arrangement on ionisation energy.[3]
- ✓ ✗ I have linked my answer to the question.[4]

Exam-style questions

Within lesson

14 **a** Ca and vanadium, phosphorus and Si

b Be and Ca

c Calcium

15 **a** 2

b [Fluorine and oxygen both have the same number of electron shells, however fluorine has nine protons whereas oxygen only has eight protons.[1]][As a result, fluorine has a greater positive charge in comparison to oxygen and is therefore able to attract valence electrons more strongly.[2]][As a result, it would take more energy to remove a valence electron from fluorine, resulting in a high ionisation energy than oxygen.[3]][Furthermore, due to the stronger attraction between fluorine's nucleus and the valence shell, fluorine would have a smaller atomic radius compared to oxygen.[4]][As a result, Element B is fluorine and Element A would represent oxygen.[5]]

- ✓ ✗ I have identified the similarity and difference between both elements.[1]
- ✓ ✗ I have compared the ability of both element to attract valence electrons.[2]
- ✓ ✗ I have compared the ionisation energy of both elements.[3]
- ✓ ✗ I have compared the atomic radius of both elements.[4]
- ✓ ✗ I have linked my answer to the question.[5]

16 **a** Sodium: +1

Potassium: +1

b [Sodium and potassium are found in the same group and also have the same core charge.[1]][However, potassium has one more energy shell than sodium, and therefore potassium would have a greater atomic radius than sodium.[2]]

- ✓ ✗ I have identified the similarity between both elements.[1]
- ✓ ✗ I have described the differences in atomic radius.[2]

c [Since potassium has a greater atomic radius than sodium, the attraction between the positive nucleus of potassium and its valence electrons is weaker than that in sodium.[1]][As a result, potassium would have a lower ionisation energy than sodium.[2]]

- ✓ ✗ I have compared the attraction of each element to its valence electrons.[1]
- ✓ ✗ I have identified the element with a lower ionisation energy.[2]

17 **a** [Fluorine and bromine are both found in Group 17 and therefore would have the same number of valence electrons and core charge.[1]] [However bromine's atomic radius is larger than fluorine due to the increase in the number of energy shells.[2]] [As a result, it would be harder to remove a valence electron from fluorine than bromine.[3]]

- ✓ ✗ I have compared the properties of both elements.[1]
- ✓ ✗ I have compared the atomic radius of both elements.[2]
- ✓ ✗ I have linked my answer to the question.[3]

b [Although both elements have the same core charge, fluorine has a smaller atomic radius compared to bromine.[1]] [As a result, fluorine is able to more strongly attract electrons to itself, and therefore would be more electronegative than bromine.[2]] [Therefore, it is most likely that fluorine has an electronegativity score of 3.98 and bromine 2.96.[3]]

- ✓ ✗ I have compared the properties of both elements.[1]
- ✓ ✗ I have compared the ability of both elements to attract electrons.[2]
- ✓ ✗ I have linked my answer to the question.[3]

c Fluorine.

Multiple lessons

18 **a** 3

b [Ionisation energy indicates how easy it is for a valence electron to be removed from an atom.[1]] [Boron contains two energy shells whereas aluminium contains three energy shells however both have the same core charge.[2]] [As a result, the attraction between the positive nucleus of aluminium and its valence electrons would be weaker than that of boron and its valence electrons.[3]] [As a result, aluminum would have a smaller ionisation energy than boron.[4]]

- ✓ ✗ I defined ionisation energy.[1]
- ✓ ✗ I have compared the similarity and difference between both elements.[2]
- ✓ ✗ I have described the difference in attraction between the nucleus and valence electrons of both elements.[3]
- ✓ ✗ I have linked my answer to the question.[4]

19 **a** Element A: 32 protons

Element B: 20 protons

b Element A: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

Element B: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

c [Both elements are found in the same period, meaning that they have the same number of energy shells, however element A has more protons in its nucleus than element B.[1]] [Therefore, the force of attraction between the positive nucleus and the valence electron of element A is greater than element B.[2]] [As a result, element A is able to pull the valence shell closer to the nucleus and therefore have a smaller atomic radius compared to element B.[3]]

- ✓ ✗ I have compared the similarity and difference between both elements.[1]
- ✓ ✗ I have compared the ability of each element to attract the valence shell.[2]
- ✓ ✗ I have linked my answer to the question.[3]

Key science skills

20 **a** Clarity of image

b **i** Qualitative

ii [Reproducibility is based on the closeness of data collected upon repeating an experiment under the same conditions.[1]] [Although test 1 and test 2 were conducted under the same conditions, the data collected varies.[2]] [Based on these results, the experiment seems to have low reproducibility.[3]]

- ✓ ✗ I have defined reproducibility.[1]
- ✓ ✗ I have identified the conditions of both experiments.[2]
- ✓ ✗ I have linked my answer to the question.[3]

2D The periodic table - part 2

Theory review questions

1 A. Highly reactive atoms are those that can gain or lose electrons easily.

2 A. Highly reactive atoms are those that can gain or lose electrons easily.

3 B. It becomes harder for elements to lose electrons as we move between groups.

4 B. Metals are known for their ability to easily lose electrons.

5 B. The most metallic elements are found in Groups 1 and 2.

6 B. Metalloids possess both metallic and non-metallic characteristics and are therefore considered to be intermediates between metals and non-metals.

7 B. Blocks in the periodic table highlight the outermost subshell.

8 B. Different blocks are found in the same period.

9 B. The p subshell can hold a maximum of six electrons.

10 B. Oxygen is found in Group 16, which is located in the p block.

Deconstructed questions

11 D **12** B

13 [Barium and magnesium are both found in Group 2.[1]] [Elements in this group are most likely to lose electrons during a reaction, therefore the easier it is for the element to lose electrons, the more reactive the element.[2]] [Barium, found in period 4, has a greater atomic radius than magnesium, found in period 2 and therefore has a weaker attraction between its nucleus and valence electrons compared to magnesium.[3]] [As a result, barium is more likely to lose electrons quicker than magnesium, and therefore would be more reactive than magnesium.[4]]

- I have identified the similarity between barium and magnesium.[1]
- I have described the reactive nature of both elements.[2]
- I have compared the atomic radius of both elements.[3]
- I have compared the reactivity of both elements.[4]

Exam-style questions

Within lesson

14 [Magnesium is found in Group 2 and is therefore an alkaline metal, whereas phosphorus is a non-metal.[1]][Metals have a distinct physical appearance in that they are shiny and silvery.[2]][As a result, sample A is most likely to contain magnesium.[3]]

- I have identified a difference between magnesium and phosphorus.[1]
- I have identified important physical characteristics of metals.[2]
- I have linked my answer to the question.[3]

15 [Given that the element in question is reactive, it cannot be Ne as it is found in Group 18 and has a full valence shell.[1]][Both C and Na are found within the first three periods and therefore cannot be the unknown element.[2]][Therefore, the unknown element is most likely to be Fr.[3]]

- I have disqualified an element based on its reactivity.[1]
- I have disqualified elements based on their position on the periodic table.[2]
- I have identified the unknown element to be Fr.[3]

16 **a** [Elements in the s block are either in Group 1 or 2, therefore are metallic elements.[1]][As a result, these elements would lose electrons during a reaction.[2]][The element with the outer 6s subshell would have the weakest attraction between the nucleus and its valence electrons, due to it having the greatest distance between the nucleus and its subshell.[3]][Therefore, the element with the outer 6s subshell would be most likely to lose electrons and is therefore the most reactive of the three elements.[4]]

- I have identified the metallic nature of the elements.[1]
- I have identified the behaviour of these elements during a reaction.[2]
- I have described the atomic structure of the 6s element.[3]
- I have established that the 6s element is the most reactive.[4]

b [Elements in the s block can be either alkali or alkaline metals found in Groups 1 and 2 respectively.[1]][These elements are the most metallic in the periodic table, therefore all three elements would have a high metallic character.[2]]

- I have identified the nature of all elements.[1]
- I have identified the metallic nature of s block elements.[2]

Multiple lessons

17 **a** $1s^2 2s^2 2p^1$

b p block

18 **a** [As the number of shells increases, the valence electrons are further from the nucleus.[1]][The shielding effect of inner shell electrons affects the ability of the nucleus to attract valence electrons.[2]][As a result, an increase in the number of electron shells increases the atomic radius.[3]]

- I have described the effect of an increase in electron shells.[1]
- I have described the effect of electron shielding on attraction between the nucleus and valence electrons.[2]
- I have explained how increasing the number of electron shells enlarges the atomic radius.[3]

b **i** Period 4

ii [The atomic radius of elements across a period decreases.[1]][This is due to the increase in the number of protons in the nucleus, resulting in an increase in the strength of the pull of valence electrons towards the nucleus.[2]]

- I have identified the trend.[1]
- I have described the impact of an increase in the number of protons on the atomic radius.[2]

c [Since both elements are found in Group 17, they both have seven valence electrons.[1]][As a result, both atoms would gain electrons in order to fill their valence shell. Therefore, the element that is more electronegative would be more reactive.[2]][Since bromine has a greater atomic radius than chlorine, it is unable to attract electrons as strongly as a chlorine atom which means it is both less electronegative and less reactive than chlorine.[3]]

- I have identified the similarity of the atomic arrangement of both elements.[1]
- I have described the property underlying the reactivity of both elements.[2]
- I have compared the reactivity of both elements.[3]

19

Statement	Correct/ Incorrect	Justification
Atoms with high reactivity are expected to have high electronegativity.	Incorrect	Metals with high reactivity lose electrons easily and have low electronegativity. Non-metals with high reactivity gain electrons easily and have high electronegativity.
Metals are more reactive than non-metals.	Incorrect	Depending on the atomic structure, elements can either gain or lose electrons during a reaction. Metals down the bottom of a group have electrons that are less attracted to the nucleus and can therefore lose electrons easily and are highly reactive. Non metals at the top of a group will gain electrons easily and are highly reactive.
Metals can be found on the left side of the periodic table.	Correct	The elements on the left side of the periodic table have few valence electrons, and so prefer to lose electrons and act as metals to satisfy the octet rule.
First ionisation energy is highest for the noble gases.	Correct	The amount of energy it takes to remove a valence electron from an atom is the first ionisation energy. Given that the noble gases have full valence shells, the amount of energy needed to remove a valence shell electron is much higher compared to the other elements.

Key science skills

20 a i Compound

ii Na: Period 3, Group 1
Cl: Period 3, Group 17

b [The data indicates that the first ionisation energy increases left-to-right across a period.[1]][This is explained by the increase in core charge across the period in the data.[2]][Given that we are comparing elements in the same period, as the core charge increases, the energy required to remove an electron increases, as they are more strongly attracted to the nucleus.[3]]

✓ ✗ I have interpreted the data.[1]

✓ ✗ I have linked core charge to the data.[2]

✓ ✗ I have explained the link between core charge and first ionisation energy.[3]

c The dependent variable is affected by a change in the independent variable – the sample. Therefore, the dependent variable is the first ionisation energy.

d [Systematic error.[1]][A systematic error affects all measurements to the same extent, in the same direction.[2]]

✓ ✗ I have identified the type of error.[1]

✓ ✗ I have explained how this error would affect the results.[2]

2E Relative masses

Theory review questions

1 B. Amu stands for atomic mass unit, and is used to express the relative atomic mass of an element.

2 B. Relative atomic mass of an element takes into consideration the different forms in which the element can exist.

3 A. ^{14}Z is the most abundant of the two isotopes, therefore the relative atomic mass is likely to be close to 14.

4 B. Each peak in a mass spectrum represents an isotope.

5 B. The least abundant isotope has the smallest peak.

6 A. The ^{90}Zr isotope is the most abundant isotope, therefore it's likely that the relative atomic mass of Zr is closer to 90.

7 B. The relative molecular mass depends on the type and number of each atom in the molecule.

8 B. The mass of a molecule depends on the relative atomic mass of all atoms found in the molecule.

9 A. Relative molecular mass is represented by M_r.

10 B. Cl_2 has two chlorine atoms in the molecule, therefore the M_r would be twice as much as the relative atomic mass of chlorine.

Deconstructed questions

11 D **12** C

13 The relative atomic mass of the element is:

$$A_r = \frac{(35 \times 75) + (37 \times 25)}{100}$$

$A_r = 35.5$

[According to the periodic table, the element with a relative atomic mass of 35.5 is chlorine.[1]]

✓ ✗ I have calculated the relative atomic mass of the element.

✓ ✗ I have then identified the name of the element.[1]

Exam-style questions

Within lesson

14 a 100 – (90.92 + 8.82) = 0.26%

b $A_r = \frac{(90.92 \times 19.99) + (0.26 \times 20.99) + (8.82 \times 21.99)}{100}$

$= 20.17$

15 a $A_r = \frac{(79 \times 24) + (10 \times 25) + (11 \times 26)}{100}$

$= 24.32$

b [The relative atomic mass of 24.32 is closest to the relative atomic mass of magnesium.[1]][Therefore, the element is most likely to be magnesium.[2]]

✓ ✗ I have used the data to determine the most likely element.[1]

✓ ✗ I have identified the unknown element.[2]

16 **a** Three

b Total peak height = 90.9 + 0.3 mm + 8.8 mm = 100 mm

Relative abundance of peak 20 = $\frac{90.9}{100} \times 100 = 90.9\%$

Relative abundance of peak 21 = $\frac{0.30}{100} \times 100 = 0.30\%$

Relative abundance of peak 22 = $\frac{8.8}{100} \times 100 = 8.8\%$

c $A_r = \frac{(90.9 \times 20) + (0.30 \times 21) + (8.88 \times 23)}{100}$

$= 20.18$

[The relative atomic mass of 20.18 is closest to the relative atomic mass of neon.[1]][Therefore, the element is most likely to be neon.[2]]

✓ ✗ I have used the data to determine the most likely element.[1]

✓ ✗ I have identified the unknown element.[2]

17 **a** ^{175}Lu

b Relative abundance of ^{175}Lu:

$174.97 = \frac{(x \times 174.94) + ((100 - x) \times 175.94)}{100}$

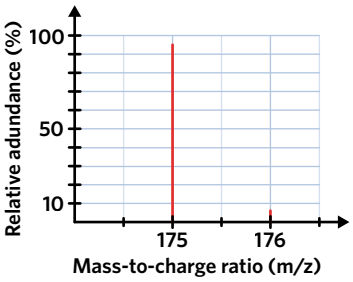
$174.97 = \frac{174.94x + 17594 - 175.94x}{100}$

$17497 = -x + 17594$

$x = 97.0\%$

Relative abundance of ^{176}Lu: 100 − 97.0 = 3.0%

c



✓ ✗ I have labelled both axes.

✓ ✗ I have drawn two peaks at 175 and 176.

✓ ✗ I have drawn the length of each peak to the abundance of each isotope.

d $M_r = 174.97 + 3 \times 72.6 = 392.77$

$= 392.8$

Multiple lessons

18 **a** S block

b $5s^1$

c $A_r = \frac{(72 \times 85) + (28 \times 87)}{100}$

$= 85.56$

$= 86$

19 **a** 29 protons

b 35 + 29 = 64

c Relative abundance of ^{63}Cu:

$63.5 = \frac{(x \times 63) + ((100 - x) \times 65)}{100}$

$63.5 = \frac{63x + 6500 - 65x}{100}$

$6350 = -2x + 6500$

$-150 = -2x$

$x = 75\%$

Relative abundance of ^{65}Cu:

100 − 75 = 25%

d Relative formula mass = 63.5 + 16.0 = 79.5

Key science skills

20 **a** [Repeatability of an experiment depends on the closeness of results measured from an experiment in identical conditions.[1]][Test 1 and test 2 were conducted by the same experimenter using the same method one after the other.[2]][The large variability in the data suggests a low repeatability.[3]]

✓ ✗ I have defined repeatability.[1]

✓ ✗ I have identified the commonality between test 1 and 2.[2]

✓ ✗ I have described the repeatability of the experiment.[3]

b **i** Three

ii Relative abundance of isotope 71 = $\frac{24.9 + 24.5}{2} = 24.7\%$

Relative abundance of isotope 73 = $\frac{75.1 + 75.5}{2} = 75.3\%$

2F The mole - part 1

Theory review questions

1 B. The number of particles in a mole is equivalent to the number of carbon atoms in 12 g of ^{12}C.

2 A. Avogadro's number is equivalent to the number of particles in one mole.

3 B. Avogadro's number is 6.02×10^{23}.

4 B. One mole of substance contains a fixed number of particles.

5 B. Avogadro's number is equivalent to one mole of substance.

6 B. The number of particles is proportional to the magnitude of change of the number of moles.

7 B. One mole will always contain a constant number of particles.

8 A. Molecules can have different compositions of elements and the number of atoms of each element.

9 A. There are two atoms of fluorine for every molecule of fluorine gas.

10 B. Each molecule contains a different number of hydrogen atoms.

Deconstructed questions

11 A **12** B

13 For every carbon dioxide molecule, there are 2 oxygen atoms produced. Therefore, in 1.22 mol of carbon, there would be twice as many moles of oxygen produced.

$n(O) = n(CO_2) \times 2$

$= 1.22 \times 2$

$= 2.44$ mol

$N(O) = 6.02 \times 10^{23} \times 2.44 = 1.47 \times 10^{24}$ atoms of oxygen

Exam-style questions

Within lesson

14 a i $n(O_2) = \frac{4.11 \times 10^{21}}{6.02 \times 10^{23}} = 6.83 \times 10^{-3}$ mol

ii $n(H_2S) = \frac{7.05 \times 10^{11}}{6.02 \times 10^{23}} = 1.17 \times 10^{-12}$ mol

iii $n(Ar) = \frac{1.28 \times 10^{2}}{6.02 \times 10^{23}} = 2.13 \times 10^{-22}$ mol

b i $N(H_2S) = 3.44 \times 6.02 \times 10^{23} = 2.07 \times 10^{24}$ molecules

ii $N(NO_3) = 0.966 \times 6.02 \times 10^{23} = 5.82 \times 10^{23}$ molecules

iii $N(SF_3) = 1.01 \times 3 \times 6.02 \times 10^{23} = 1.82 \times 10^{24}$ atoms of fluorine

15 N(hydrogen) $= 4.78 \times 6.02 \times 10^{23} = 2.88 \times 10^{24}$ molecules

[Of the two methods, the method requiring 4.78 mol of hydrogen requires 2.88×10^{24} molecules of hydrogen compared to the second method, which requires 5.63×10^{16} molecules.[1]][As a result, the method requiring 4.78 mol of hydrogen would use more hydrogen gas.[2]]

- ✓ ✗ I have included my calculations and compared the quantity of hydrogen required for both methods.[1]
- ✓ ✗ I have identified the method requiring more hydrogen.[2]

16 a 2.11 mol of iron (Fe) atoms < 3.33×10^{33} molecules of chlorine gas (Cl_2).

b 5.17×10^{11} molecules of nitrogen dioxide (NO_2) < 2.98 mol of carbon tetrachloride (CCl_4).

c 7.01×10^{9} molecules of carbon disulfide (CS_2) < hydrogen atoms in 2.70×10^{9} molecules of phosphine (PH_3).

Multiple lessons

17 a $M_r(FeO) = (55.8 + 16.0) = 71.8$

b N(O atoms) $= 1.45 \times 2 \times 6.02 \times 10^{23} = 1.75 \times 10^{24}$ atoms of oxygen

18 a 18

b $1s^2 2s^2 2p^6 3s^2 3p^6$

c [The relative atomic mass is derived from the average mass of all isotopes, taking into account the relative abundance of each isotope.[1]][According to the periodic table, the relative atomic mass of argon is 39.9.[2]][This value is closest to the value of the mass of ^{40}Ar, which is therefore the most abundant isotope.[3]]

- ✓ ✗ I have described how the relative atomic mass is calculated.[1]
- ✓ ✗ I have identified the relative atomic mass of Ar.[2]
- ✓ ✗ I have identified the most abundant isotope.[3]

d N(argon) $= 2.44 \times 6.02 \times 10^{23} = 1.47 \times 10^{24}$ atoms

19 a Both are in period 2.

b [Both nitrogen and oxygen contain two energy shells.[1]][However, oxygen has one more proton in its nucleus than nitrogen and as a result would have a greater core charge. Consequently, it has a greater ability to attract electrons to itself.[2]][Since electronegativity describes the ability of an atom to attract electrons, oxygen is therefore more electronegative than nitrogen.[3]]

- ✓ ✗ I have identified the similarity between both elements.[1]
- ✓ ✗ I have explained the difference in each element's ability to attract electrons.[2]
- ✓ ✗ I have identified the element that is more electronegative.[3]

c [Reactivity describes an atom's tendency to either lose or gain electrons.[1]][Since oxygen has a greater tendency to gain electrons compared to nitrogen, oxygen would be more reactive than nitrogen.[2]]

- ✓ ✗ I have defined reactivity.[1]
- ✓ ✗ I have compared the behaviour of both elements and identified the more reactive element.[2]

d $n(N_2O) = \frac{2.56 \times 10^{24}}{6.02 \times 10^{23}} = 4.25$ mol

Key science skills

20 a The amount of methane collected.

b N(methane) $= 15.8 \times 6.02 \times 10^{23} = 9.51 \times 10^{24}$ molecules

c [One of the main environmental concerns with this experiment is that the miners need to drill and dig a deep hole into the earth.[1]][As a result, this could destroy the land in the area being dug, and/or disturb the organisms living in the area.[2]]

- ✓ ✗ I have identified the key environmental concern.[1]
- ✓ ✗ I have explained the effect that drilling and digging would have on the environment.[2]

2G The mole - part 2

Theory review questions

1 A. The molar mass is the mass of one mole of substance.

2 B. Different types of atoms have varying masses.

3 B. The number that represents the relative atomic mass of an element is the same number (with different units) as the element's molar mass.

4 A. The molar mass is the total mass of a mole of the substance.

5 B. We need to know both the molar mass and mass of a substance to find the amount.

6 B. The mass of a substance depends on the amount available.

7 B. The amount of substance indicates the number of 'groups' of the molar mass found in the substance.

8 A. The molar mass is found by dividing the mass of substance by the amount.

9 B. Both equations calculate the amount of substance.

10 A. A mole of substance contains a fixed number of particles.

Deconstructed questions

11 D **12** D

13 $n(NH_3) = \frac{6.11 \times 10^{26}}{6.02 \times 10^{23}} = 1014.95$ mol

$m(NH_3) = 1014.95 \times (14.0 + 3 \times 1.0) = 17254.2$ g $= 1.73 \times 10^4$ g

Exam-style questions

Within lesson

14 a i $M(HI) = 1.0 + 126.9 = 127.9\ g\ mol^{-1}$

ii $M(MgCl_2) = 24.3 + (2 \times 35.5) = 95.3\ g\ mol^{-1}$

iii $M(C_3H_6O_3) = (3 \times 12.0) + (6 \times 1.0) + (3 \times 16.0) = 90.0\ g\ mol^{-1}$

b i $NiCl_2$ > HNO_3

ii H_3PO_4 > KCl

iii KCNS > HCOONa

15 a i $n(SF_3) = \frac{14.9}{(32.1 + (19.0 \times 3))} = 0.167\ mol$

ii $n(C_3H_6O) = \frac{0.552}{(3 \times 12.0) + (6 \times 1.0) + 16.0} = 9.52 \times 10^{-3}\ mol$

iii $n(KCN) = \frac{1.12}{39.1 + 12.0 + 14.0} = 1.72 \times 10^{-2}\ mol$

b i $n(H_2) = \frac{4.60}{(2 \times 1.0)} = 2.3\ mol$

$N(H_2) = 2.3 \times 6.02 \times 10^{23} = 1.38 \times 10^{24}$ molecules

ii $n(O_2) = \frac{2.10}{(2 \times 16.0)} = 0.06563\ mol$

$N(O_2) = 0.06563 \times 6.02 \times 10^{23} = 3.95 \times 10^{22}$

$N(O) = 2 \times 3.951 \times 10^{22} = 7.90 \times 10^{22}$ atoms of oxygen

iii $n(Be) = \frac{390{,}531}{6.02 \times 10^{23}} = 6.487 \times 10^{-19}\ mol$

$m(Be) = 6.487 \times 10^{-19} \times 9.0 = 5.84 \times 10^{-18}\ g$

16 a $n(C_6H_{12}O_6) = \frac{4.11}{(6 \times 12.0) + (12 \times 1.0) + (6 \times 16.0)} = 2.28 \times 10^{-2}\ mol$

b $M(C_6H_{12}O_6) = (6 \times 12.0) + (12 \times 1.0) + (6 \times 16.0) = 180.0\ g\ mol^{-1}$

$M(C_2H_5OH) = (2 \times 12.0) + (6 \times 1.0) + 16.0 = 46.0\ g\ mol^{-1}$

[Glucose, with a molar mass of 180.0 g mol^{-1} has a higher molar mass than ethanol, which has a molar mass of 46.0 g mol^{-1}.[1]]

✓ ✗ I have included calculations in my answer.

✓ ✗ I have used data to identify the substance with a greater molar mass.[1]

c $n(C_6H_{12}O_6)_{required} = 3.00 - 0.0228 = 2.977\ mol$

$m(C_6H_{12}O_6)_{required} = 2.977 \times ((6 \times 12) + (12 \times 1.0) + (6 \times 16)) = 535.9 = 536\ g$

Multiple lessons

17 a $m(CO) = 201.3 \times (12.0 + 16.0) = 5636.4 = \frac{5636.4}{1000} = 5.64\ kg$

b Car A: $n(CO) = \frac{114}{(12.0 + 16.0)} = 4.07\ mol$

Car B: $n(CO) = \frac{12.04 \times 10^{33}}{6.02 \times 10^{23}} = 2.00 \times 10^{10}\ mol$

[In comparison to Car B, which releases 2.00×10^{10} mol of CO, Car A releases less CO, at only 4.07 mol per 40L tank of fuel.[1]]

✓ ✗ I have included calculations in my answer.

✓ ✗ I have identified the car that produces the least amount of CO.[1]

18 a 

b $1s^2 2s^2 2p^2$

c $m(C) = 17.41 - 10 - 5 = 2.41\ g$

$n(C) = \frac{2.41}{12.0} = 0.201\ mol$

19 a $m(C_6H_{12}O_6) = 0.98 \times ((6 \times 12.0) + (12 \times 1.0) + (6 \times 16.0)) = 176.4\ g = 176\ g$

b $m(NaHCO_3) = 0.11 \times (23.0 + 1.0 + 12.0 + (3 \times 16.0)) = 9.24\ g$

Mass of both = 9.24 + 176.4 = 185.6 g

c Sodium, period 3

Key science skills

20 a Size of the chocolate

b [Reliability describes the consistency of results of the same experiment conducted in different environments.[1]][Due to the fact that the type of chocolate used was changed, this introduced errors that could affect the results of the experiment. For example, the samples would have differed in their sugar, fat and cocoa content, all of which could affect the rate of melting.[2]][As a result, it reduces the reliability of the results.[3]]

✓ ✗ I have defined reliability.[1]

✓ ✗ I have identified the effect of the change in chocolate.[2]

✓ ✗ I have identified the effect of the change on the reliability of the results.[3]

2H Chemical formulas and equations

Theory review questions

1 A. The compound only contains carbon and hydrogen and therefore, the percentage by mass of carbon and hydrogen will add up to 100%.

2 B. Percentage by mass of an element is calculated by the mass of the element in 1 mol of the compound divided by the molar mass of the compound.

3 A. Molecular formulas depict the actual number of atoms of each element present.

4 B. Empirical formulas depict the simplest, whole-number ratio of atoms of each element present.

5 B. Sometimes the molecular formula is the same as the empirical formula.

6 B. The empirical formula can be determined from the molecular formula by identifying a common factor to simplify the ratio.

7 A. The molar mass of the compound is required to determine the number of empirical formula units in the molecular formula.

8 B. There is one sodium atom on the left-hand side of the equation compared to two on the right-hand side.

9 B. There are two atoms of aluminum in Al_2O_3 and two atoms of iron in Fe_2O_3.

10 A. The numbers in front of the molecular formulas represent the number of molecules of the compound present.

Deconstructed questions

11 C **12** D

13

Fe	:	O
0.00760 mol	:	0.0114 mol
$\frac{0.00760}{0.00760}$	:	$\frac{0.0114}{0.00760}$
1.00	:	1.50

Multiply by 2 to get the simplest whole number ratio:

2.00	:	3.00

Therefore, the empirical formula is Fe_2O_3.

[The molar mass of the empirical formula is $2 \times 55.8 + 3 \times 16.0 = 159.6$ g mol^{-1}.[1]] [Since the molar mass of the empirical formula is the same as the molar mass of the compound's molecular formula, the empirical formula and molecular formula are the same.[2]]

✓ ✗ I have included calculations.

✓ ✗ I have calculated the molar mass of the empirical formula.[1]

✓ ✗ I have determined the molecular formula.[2]

Exam style questions

Within lesson

14 a Mass of H in 1 mol of HCl = 1.0 g

$M(HCl) = (1 \times 1.0) + (1 \times 35.5) = 36.5$ g mol^{-1}

$\%\ mass\ by\ composition = \frac{1.0}{36.5} \times 100 = 2.7\%$

b Mass of K in 1 mol of $KMnO_4$ = 39.0 g

$M(KMnO_4) = (1 \times 39.0) + (1 \times 54.9) + (4 \times 16.0) = 157.9$ g mol^{-1}

$\%\ mass\ by\ composition = \frac{39.0}{157.9} \times 100 = 24.7\%$

c Mass of Mn in 1 mol of $KMnO_4$ = 54.9 g

$M(KMnO_4) = (1 \times 39.0) + (1 \times 54.9) + (4 \times 16.0) = 157.9$ g mol^{-1}

$\%\ mass\ by\ composition = \frac{54.9}{157.9} \times 100 = 34.8\%$

d Mass of N in 1 mol of $(NH_4)_3PO_4 = 3 \times 14.0 = 42.0$ g

$M((NH_4)_3PO_4) = (3 \times 14.0) + 3 \times (4 \times 1.0) + (1 \times 31.0) + (4 \times 16.0) = 149.0$ g mol^{-1}

$\%\ mass\ by\ composition = \frac{42.0}{149.0} \times 100 = 28.2\%$

e Mass of H in 1 mol of $(NH_4)_3PO_4 = 3 \times 4 \times 1.0 = 12.0$ g

$M((NH_4)_3PO_4) = (3 \times 14.0) + 3 \times (4 \times 1.0) + (1 \times 31.0) + (4 \times 16.0) = 149.0$ g mol^{-1}

$\%\ mass\ by\ composition = \frac{12.0}{149.0} \times 100 = 8.05\%$

f Mass of Fe in 1 mol of $Fe_2O_3 = 2 \times 55.8 = 111.6$ g

$M(Fe_2O_3) = (2 \times 55.8) + (3 \times 16.0) = 159.6$ g mol^{-1}

$\%\ mass\ by\ composition = \frac{111.6}{159.6} \times 100 = 69.92\%$

$m(Fe) = 0.6992 \times 36.2 = 25.3$ g

15 a

H	:	Cl
2.8%	:	97.2%
2.8 g	:	97.2 g
$\frac{2.8}{1.0}$	:	$\frac{97.2}{35.5}$
2.8 mol	:	2.74 mol
$\frac{2.8}{2.74}$	:	$\frac{2.74}{2.74}$
1.00	:	1.00

Empirical formula: HCl

b

Pb	:	C	:	H
38.43 g	:	17.83 g	:	3.74 g
$\frac{38.43}{207.2} = 0.185$	:	$\frac{17.83}{12.0} = 1.49$ mol	:	$\frac{3.74}{1.0} = 3.74$ mol
$\frac{0.185}{0.185}$	:	$\frac{1.49}{0.185}$	:	$\frac{3.74}{0.185}$
1.00	:	8.05	:	20.2

Empirical formula: PbC_8H_{20}

c

Al	:	S	:	O
15.8%	:	28.1%	:	56.1%
15.8 g	:	28.1 g	:	56.1 g
$\frac{15.8}{27.0}$	:	$\frac{28.1}{32.1}$	:	$\frac{56.1}{16.0}$
0.585 mol	:	0.875 mol	:	3.51 mol
$\frac{0.585}{0.585}$	:	$\frac{0.875}{0.585}$	:	$\frac{3.51}{0.585}$
1.00	:	1.50	:	6.00
2.00	:	3.00	:	12.00

Empirical formula: $Al_2S_3O_{12}$

d

C	:	H
4.8 g	:	6.4 g – 4.8 g
4.8 g	:	1.6 g
$\frac{4.8}{12.0}$	:	$\frac{1.6}{1.0}$
0.40 mol	:	1.6 mol
$\frac{0.40}{0.40}$	:	$\frac{1.6}{0.40}$
1.0	:	4.0

Empirical formula: CH_4

e

N	:	O
0.608 g	:	1.996 g – 0.608 g
0.608 g	:	1.388 g
$\frac{0.608}{14.0}$	:	$\frac{1.388}{16.0}$
0.0434 mol	:	0.0868 mol
$\frac{0.0434}{0.0434}$	:	$\frac{0.0868}{0.0434}$
1.00	:	2.00

Empirical formula: NO_2

16 a Molar mass of empirical formula = $(1 \times 12.0) + (2 \times 1.0) = 14.0$ g mol^{-1}

Empirical formula units = $\frac{70.0}{14.0} = 5.0$

Molecular formula is $C_{1 \times 5}H_{2 \times 5} = C_5H_{10}$

b Molar mass of empirical formula = (1 × 12.0) + (2 × 1.0) + (1 × 16.0) = 30.0 g mol^{-1}

Empirical formula units = $\frac{180.0}{30.0}$ = 6.00

Molecular formula is $C_{1\times6}H_{2\times6}O_{1\times6} = C_6H_{12}O_6$

c Molar mass of empirical formula = (6 × 12.0) + (10 × 1.0) + (2 × 32.1) + (1 × 16.0) = 162.2 g mol^{-1}

Empirical formula units = $\frac{162.2}{162.2}$ = 1.00

Molecular formula is $C_{6\times1}H_{10\times1}S_{2\times1}O_{1\times1} = C_6H_{10}S_2O$

d Molar mass of empirical formula = (1 × 12.0) + (4 × 1.0) + (1 × 14.0) = 30.0 g mol^{-1}

Empirical formula units = $\frac{60.0}{30.0}$ = 2.00

Molecular formula is $C_{1\times2}H_{4\times2}N_{1\times2} = C_2H_8N_2$

e Molar mass of empirical formula = (3 × 12.0) + (8 × 1.0) + (1 × 14.0) = 58.0 g mol^{-1}

Empirical formula units = $\frac{116.0}{58.0}$ = 2.00

Molecular formula is $C_{3\times2}H_{8\times2}N_{1\times2} = C_6H_{16}N_2$

17 a i $CH_{4(g)} + 4Cl_{2(g)} \rightarrow CCl_{4(g)} + 4HCl_{(g)}$

ii One mole of $CH_{4(g)}$ reacts with four moles of $Cl_{2(g)}$ to produce one mole of $CCl_{4(g)}$ and four moles of $HCl_{(g)}$.

b i $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

ii Two moles of $Mg_{(s)}$ reacts with one mole of $O_{2(g)}$ to produce two moles of $MgO_{(s)}$.

iii $n(Mg) = \frac{2}{2} \times = 1$ mol

$n(O_2) = \frac{1}{2} \times 1 = 0.5$ mol

c i $2Al_{(s)} + 3Cl_{2(g)} \rightarrow 2AlCl_{3(g)}$

ii Two moles of $Al_{(s)}$ reacts with three moles of $Cl_{2(g)}$ to produce two moles of $AlCl_{3(g)}$.

iii $n(Al) = \frac{2}{2} \times 1 = 1$ mol

$n(Cl_2) = \frac{3}{2} \times 1 = 1.5$ mol

d i $SeCl_{6(g)} + O_{2(g)} \rightarrow SeO_{2(s)} + 3Cl_{2(g)}$

ii One mole of $SeCl_{6(g)}$ reacts with one mole of $O_{2(g)}$ to produce one mole of $SeO_{2(s)}$ and three mole of $Cl_{2(g)}$.

e i $4H_2O_{(l)} + 7CO_{2(g)} \rightarrow C_7H_{8(l)} + 9O_{2(g)}$

ii Four moles of $H_2O_{(l)}$ reacts with seven moles of $CO_{2(g)}$ to produce one mole of $C_7H_{8(l)}$ and nine moles of $O_{2(g)}$.

Multiple lessons

18 a $Li_{(s)} + H_2O_{(l)} \rightarrow LiOH_{(aq)} + H_{2(g)}$

b $2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$

[The above equations are now balanced because there is the same number of lithium, hydrogen and oxygen atoms on both sides of the chemical equation.[1]]

✓ ✗ I have written the balanced chemical equation.

✓ ✗ I have justified why the chemical equation is balanced.[1]

c $1s^22s^1$

d [Lithium and fluorine are in the same period of the periodic table.[1]] [However, fluorine is is more electronegative (ability to attract an electron) because it has a greater core charge with an increased number of protons and same number of inner shell electrons compared to lithium.[2]][This allows fluorine to attract electrons more strongly than lithium.[3]]

✓ ✗ I have identified the commonality between lithium and fluorine.[1]

✓ ✗ I have justified my answer with reference to the trends in the periodic table.[2]

✓ ✗ I have compared the ability of both elements to attract electrons.[3]

e [Lithium and potassium are in the same group of the periodic table and also have the same core charge.[1]][However, potassium has three electron shells compared to lithium's two, resulting in the valence electron of potassium being further from the nucleus.[2]][Therefore, the attraction between the valence electron and the nucleus of potassium is weaker, which results in potassium being able to lose its valence electron easier than lithium.[3]][As a result, potassium is more reactive than lithium.[4]]

✓ ✗ I have identified the commonalities between lithium and potassium.[1]

✓ ✗ I have compared the structure of lithium and potassium.[2]

✓ ✗ I have described the effect of the number of electron shells on the ability of the atoms to lose electrons.[3]

✓ ✗ I have linked my answer to the question.[4]

19 a $n(H_2O) = \frac{0.255}{18.0} = 0.0142$ mol

$n(H)$ *in compound* $= 2 \times n(H_2O) = 0.0284$ mol

$n(CO_2) = \frac{1.24}{44.0} = 0.0282$ mol

$n(C)$ *in compound* $= 1 \times n(CO_2) = 0.0282$ mol

C : H

0.0282 mol : 0.0284 mol

$\frac{0.0282}{0.0282} : \frac{0.0284}{0.0282}$

1.00 : 1.01

Therefore, the empirical formula is CH.

b Molar mass of empirical formula = (1 × 12.0) + (1 × 1.0) = 13.0 g mol^{-1}

Empirical formula units = $\frac{78.0}{13.0}$ = 6.0

Molecular formula is $C_{1\times6}H_{1\times6} = C_6H_6$

c $C_6H_{6(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

d $2C_6H_{6(l)} + 15O_{2(g)} \rightarrow 12CO_{2(g)} + 6H_2O_{(g)}$

[The equation is now balanced because there are the same number of carbon, hydrogen and oxygen atoms on either side of the chemical equation.[1]]

✓ ✗ I have written the balanced chemical equation.

✓ ✗ I have justified why the chemical equation is balanced.[1]

e [Carbon and oxygen are in the same period of the periodic table.[1]] [However, oxygen is more electronegative because it has a greater core charge due to an increased number of protons in the nucleus and the same number of inner-shell electrons as carbon, resulting in a greater ability to attract electrons to itself.[2]]

✓ ✗ I have identified the relationship between carbon and oxygen.[1]

✓ ✗ I have justified my answer with reference to the trends in the periodic table.[2]

Key science skills

20 a [The carbon dioxide measurement will be accurate but not the water measurement.[1]] [This is because accuracy is a measure of how close the measured value is to the true value, and since the water absorber was exposed to the atmosphere, atmospheric water would have been absorbed, reducing the accuracy of the results.[2]]

✓ ✗ I have commented on the accuracy of the results.[1]

✓ ✗ I have justified my answer with reference to the definition of accuracy.[2]

b [It is a systematic error.[1]] [This is because the uncalibrated scale will provide measurements which are always incorrect by the same degree.[2]]

✓ ✗ I have identified the type of error.[1]

✓ ✗ I have justified my answer with reference to the definition of a systematic error.[2]

c [The experimenter is incorrect.[1]] [This is because the results will still be precise, since the values are all affected to the same extent and therefore, will be close to each other. However, the results will be inaccurate because the measurements are not the true values as the scale is uncalibrated.[2]]

✓ ✗ I have commented on the experimenter's statement.[1]

✓ ✗ I have justified my answer with reference to the definition of precision and accuracy.[2]

Chapter 2 review

Multiple choice questions

1 D. Bohr's model of the atom describes electrons circling around the nucleus of an atom in fixed orbitals of specific energy levels.

2 A. Period 3 indicates that there are three electron shells and Group II indicates that there are two valence electrons.

3 C. The atomic number represents the number of protons in the nucleus of an atom which may be different to the number of neutrons.

4 D. The mass number is equal to the sum of the number of protons (atomic number) and the number of neutrons.

5 D. Metallic character decreases from left to right (more difficult to remove electrons) and increases down a group (easier to remove electrons).

6 C. The empirical formula is the simplest whole number ratio of atoms of each element in a molecule.

7 A. Selenium has a relative atomic mass of 79.0.

8 D. There are three moles of oxygen per mole of $CaCO_3$.

9 C. There are six moles of oxygen per mole of $Ba(NO_3)_2$.

10 B. Balanced equations have the same number of atoms of each element on either side of the equation.

Short answer questions

11 a Name: Titanium

Chemical symbol: Ti

Atomic number: 22

Relative atomic mass: 47.9

b [Titanium is found in the d-block.[1]] [This is because its highest energy electrons are in the d-subshell.[2]]

✓ ✗ I have identified which block on the periodic table the element is found in.[1]

✓ ✗ I have justified my answer with reference to its electron configuration.[2]

c [Electronegativity is a measure of how strongly electrons are attracted to the nucleus of an atom.[1]] [Electronegativity increases across a period from left to right as the core charge increases and therefore, titanium will be less electronegative than copper.[2]]

✓ ✗ I have defined electronegativity.[1]

✓ ✗ I have compared the electronegativities of copper and the element.[2]

d [Metallic character is based on an element's ability to lose electrons easily.[1]] [Since zirconium is further down in the group, it has the same core charge but its valence electrons are further away, resulting in greater metallic character as it can lose electrons more easily than titanium.[2]]

✓ ✗ I have defined metallic character.[1]

✓ ✗ I have compared the metallic character of copper and zirconium.[2]

e [Titanium is a metal.[1]] [Therefore, it is more likely to react by losing electrons as opposed to gaining electrons.[2]]

✓ ✗ I have identified titanium as a metal.[1]

✓ ✗ I have identified its preferred mechanism of reactivity.[2]

12 a $n(H_2O) = \frac{3.21}{18.0} = 0.178$ mol

$n(H) = 2 \times n(H_2O) = 0.357$ mol

$m(H)$ *in compound* $= 0.357 \times M(H) = 0.357 \times 1.0 = 0.36$ g

$n(CO_2) = \frac{7.85}{44.0} = 0.178$ mol

$n(C) = 1 \times n(CO_2) = 0.178$ mol

$m(C)$ *in compound* $= 0.178 \times M(C) = 0.178 \times 12.0 = 2.14$ g

$m(O)$ *in compound* = *total mass of compound* − $m(C) - m(H)$

$= 3.07 - 2.14 - 0.36 = 0.57$ g

$n(O) = \frac{0.57}{16.0} = 0.036$ mol

REVIEW

C	:	H	:	O
0.178 mol	:	0.357 mol	:	0.036 mol
$\frac{0.178}{0.036}$	:	$\frac{0.357}{0.036}$	:	$\frac{0.036}{0.036}$
4.9	:	9.9	:	1.0

Therefore, the empirical formula is $C_5H_{10}O$.

b Molar mass of empirical formula
$= (5 \times 12.0) + (10 \times 1.0) + (1 \times 16.0) = 86.0\ g\ mol^{-1}$

Empirical formula units $= \frac{172.0}{86.0} = 2$

Molecular formula is $2 \times (C_5H_{10}O) = C_{5\times2}H_{10\times2}O_{1\times2} = C_{10}H_{20}O_2$

c $C_{10}H_{20}O_{2(l)} + 14O_{2(g)} \rightarrow 10CO_{2(g)} + 10H_2O_{(g)}$

d An isotope is an atom of the same element (same number of protons) but with a different number of neutrons.

e [The one small peak corresponds to a molecule with one ^{13}C atom instead of all ^{12}C atoms.[1]][This is because the molecular mass is one greater than what is expected and the ^{13}C isotope is the most common compared to ^{2}H or ^{17}O.[2]]

✓ ✗ I have identified what causes the peak at m/z 173.[1]

✓ ✗ I have justified my answer with reference to the relative abundances of isotopes.[2]

f [Nanoparticles are spherical in shape and are measured on the nano (10^{-9}) scale.[1]][Due to their small size, they are able to be absorbed by the body and can remain indefinitely in the body.[2]]

✓ ✗ I have defined nanoparticles.[1]

✓ ✗ I have explained why health concerns are being raised about nanoparticles.[2]

13 a [^{63}Cu is more abundant in nature.[1]][This is because the relative atomic mass of copper is closer to 63 than it is to 65, indicating that there is a greater percentage of ^{63}Cu isotopes present than ^{65}Cu isotopes.[2]]

✓ ✗ I have specified which isotope is more abundant in nature.[1]

✓ ✗ I have justified my answer.[2]

b $2Cu_2O_{(s)} \rightarrow O_{2(g)} + 4Cu_{(s)}$

c According to the balanced chemical equation, two moles of Cu_2O decompose to produce one mole of $O_{2(g)}$ and four moles of $Cu_{(s)}$.

$n(O_2) = \frac{1}{2} \times 1.00 = 0.50$ mol

$N = 6.02 \times 10^{23} \times \frac{1}{2} = 3.01 \times 10^{23}$ molecules

d [First ionisation energy is the energy required to remove the first valence electron from an atom.[1]][As we move across a period from left to right, or from copper to bromine, first ionisation energy increases because core charge increases and so the valence electrons are more strongly attracted to the nucleus of an atom.[2]]

✓ ✗ I have defined first ionisation energy.[1]

✓ ✗ I have explained the trend in the periodic table.[2]

Key science skills questions

14 a The aim of this experiment is to prove the law of conservation of mass by carrying out a reaction between mercury and oxygen in a sealed vessel.

b If the law of conservation of mass is true, there should be no change in the mass of the sealed vessel before and after the chemical reaction has occurred.

c [The mass of the sealed vessel before and after the experiment would need to be collected.[1]][This is quantitative data.[2]][The reason this data needs to be collected is to prove that even though a chemical reaction has occurred, the mass of the vessel has not changed which means that atoms have only been rearranged.[3]]

✓ ✗ I have identified what data needs to be collected.[1]

✓ ✗ I have specified what type of data it is.[2]

✓ ✗ I have justified why this data needs to be collected.[3]

d [If the jar is not sealed properly and reactant (oxygen) escapes, the recorded mass of the vessel would be lower after the reaction has occurred compared to the initial mass, apparently disproving the law of conservation of mass.[1]][This is a random error.[2]]

✓ ✗ I have identified the effect on the results the error has.[1]

✓ ✗ I have classified the type of error.[2]

e Personal protective equipment including gloves, glasses and a lab coat should be worn.

3A Metals

Theory review questions

1 B. Metals are extracted from ores which are the deposits of minerals containing metals.

2 A. The original ore sample needs to be refined before it can be treated.

3 B. CO helps to extract the metal from the ore sample.

4 B. The names of electrons indicate whether they can move freely or they are fixed.

5 B. Delocalised electrons are removed from metal atoms and therefore are attracted to metal cations.

6 A. Malleability, ductility, lustre, heat and electrical conductivity can be explained by the metallic bonding model.

7 A. Electrostatic force of attraction in the metallic bonding model is the attractive force between delocalised electrons and metal cations.

8 A. Delocalised electrons are removed from metal atoms and therefore can absorb heat and transfer kinetic energy to other delocalised electrons.

9 B. Delocalised electrons have negative charge so they are pushed by the negative electrode and attracted by the positive electrode.

10 A. Metallic bonds are strong so they can withstand high temperature.

Deconstructed questions

11 A 12 C

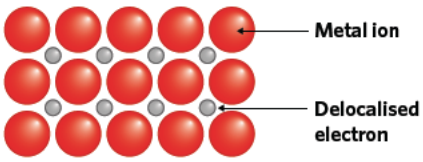
13 [Based on the data given, iron has the highest melting point of all the metals.[1]][This suggests that the metallic bonds involving iron are the strongest.[2]][Therefore, it would be able to withstand more weight and heat compared to the other metals.[3]][As a result, iron would be the most suitable metal for the company to use.[4]]

- ✓ ✗ I have identified the metal with the highest melting point.[1]
- ✓ ✗ I have identified the metal with the strongest bonds.[2]
- ✓ ✗ I have described the effect of bond strength on different properties.[3]
- ✓ ✗ I have linked my answer to the question.[4]

Exam-style questions

Within lesson

14 a



b [Metals are held together by the force of attraction that exists between the positive metal cation and the sea of negatively charged delocalised electrons.[1]][When a force is applied to the lattice structure, the layers of positive ions can move, however due to the delocalised electrons, the force of attraction between the electrons and positive ions remains,[2]][able to still hold the structure together and therefore allowing it to be bent without breaking.[3]]

- ✓ ✗ I have described metallic bonding.[1]
- ✓ ✗ I have explained what happens when a force is applied to a metallic structure.[2]
- ✓ ✗ I have linked my answer to the question.[3]

15 a [During the extraction of metals, a large amount of rocks and sand collected can destruct the landscape and therefore result in flooding, soil erosion and water pollution.[1]][Moreover, released greenhouse gases such as carbon dioxide and sulfur dioxide can lead to global warming and acid rain and the dust released decreases the air quality.[2]]

- ✓ ✗ I have described the impact of processes involved in metal extraction.[1]
- ✓ ✗ I have described the impact of the products of this process on the atmosphere.[2]

b [Due to the negative impacts of the extraction of metals on the environment, communities living near extraction sites suffer from noise, air, land and water pollution,[1]][which causes them to have many health issues such as respiratory and digestive diseases and most seriously cancer.[2]]

- ✓ ✗ I have described how the extraction of metals can impact nearby environments.[1]
- ✓ ✗ I have identified the health impacts related to this process.[2]

16 a Nick can conclude that metals can conduct heat.

b [According to the metallic bonding model, there are many delocalised electrons between metal cations.[1]][When a piece of metal is heated, the rise in temperature causes the increase in kinetic energy of delocalised electrons and metal cations. Therefore, the metallic ions can vibrate more rapidly and delocalised electrons can move faster, transferring the kinetic energy to nearby ions and electrons.[2]][Hence, when Nick touched the spoon, the kinetic energy carried by moving electrons were transferred to his hand in the form of heat and therefore he had a 'burning' feeling.[3]]

- ✓ ✗ I have described delocalised electrons in the metallic bonding model.[1]
- ✓ ✗ I have used the metallic bonding to explain how metals can conduct heat.[2]
- ✓ ✗ I have linked my explanation with Nick's experience.[3]

17 [The student's explanation is not correct.[1]][For a material to be able to conduct electricity, it needs to be able to allow for the movement of electrons.[2]][Metals can conduct electricity because according to the metallic bonding model, delocalised electrons in a solid piece of metal can move freely. Therefore, when one end of the metal is connected to the positive electrode and the other end is connected to the negative electrode of an electrical source, there is a movement of delocalised electrons away from the negative electrode and towards the positive electrode.[3]]

- ✓ ✗ I have determined if the student's explanation is correct.[1]
- ✓ ✗ I have described metal conductivity.[2]
- ✓ ✗ I have explained how metals can conduct electricity.[3]

18 a [According to the results, the structure of both metals are unaffected at 200°C. At 400°C the lead rail began to lose its structure whereas copper remained unaffected.[1]][This suggests that the energy supplied at 400°C was enough to break the metallic bonds within the lead rail, implying that the melting temperature of lead is within range of 400°C.[2]][In contrast, the structure of copper's rail remained unchanged, suggesting that the melting point for copper is greater than 600°C.[3]][Through these results, it can be seen that lead has a lower melting point than copper.[4]]

- ✓ ✗ I have identified the key results in the experiment.[1]
- ✓ ✗ I have explained the effect of heat on the bonding of lead.[2]
- ✓ ✗ I have explained the effect of heat on the bonding of copper.[3]
- ✓ ✗ I have linked my answer to the question.[4]

b [Although the metallic bonding model is able to explain most of the properties of metals, it isn't able to explain why different metals exhibit slightly different variations of certain properties.[1]][For example, for metals to conduct heat, the kinetic energy is passed on through the movement of electrons [2]][however since all metallic structures contain delocalised electrons, we are unable to use the model to explain why different metals are able to conduct heat more or less efficiently.[3]]

- ✓ ✗ I have identified the main limitation of the metallic bonding model.[1]
- ✓ ✗ I have described how heat is conducted through metals.[2]
- ✓ ✗ I have described the inability of the metallic bonding model to explain heat conductivity.[3]

Multiple lessons

19 a Titanium

b [Titanium is a transition metal and magnesium is a Group 2 metal.[1]][According to the trend in the periodic table, titanium has higher nuclear charge than magnesium, meaning that the electrons of a titanium atom are held more strongly to the nucleus. Therefore, titanium has a smaller atomic size than magnesium.[2]][This means that there is less space between titanium cations in the metallic bonding model and therefore they are packed more tightly and closely together.[3]][Hence, titanium can withstand stronger bending force without breaking than magnesium so the engineer should choose titanium.[4]]

- ✓ ✗ I have determined the most appropriate metal to be used.[1]
- ✓ ✗ I have compared the atomic size of titanium and magnesium.[2]
- ✓ ✗ I have linked the difference in atomic size to the difference in the metallic bonding model of titanium and magnesium.[3]
- ✓ ✗ I have used my explanation to justify my answer.[4]

Key science skills

20 a [Gabriella can conclude that her ring is not entirely made of silver.[1]][There was dull light from the bulb when the ring was implemented in the circuit, while there was strong and dazzling light from the bulb when the piece of pure silver was implemented in the circuit.[2]]

- ✓ ✗ I have given a conclusion.[1]
- ✓ ✗ I have used my observation in the experiment to explain my conclusion.[2]

b Gabriella can use different sets of wires which are made of different metals to confirm that the same results will be obtained.

3B Variations of metallic substances

Theory review questions

1 A. Metals can be mixed with other elements to produce alloys.

2 B. Atoms in a substitutional alloy are similar in size.

3 B. In an interstitial alloy, alloying agent atoms need to be much smaller than main metal atoms.

4 A. Atoms with smaller sizes than main metal atoms lead more rigid structures.

5 B. Slow cooling allows for bigger crystals to form.

6 A. Annealing allows heated metals to cool down for a longer time than quenching.

7 B. Coatings provide physical barriers between the surface of metals and oxygen or other damaging chemicals.

8 B. Electroplating requires an electrical current to adhere ions from a solution of cadmium and chromium to a metal, while powder coating is the process of applying a dry powder to the surface of a metal.

9 A. The lattice structure of metallic nanoparticles is very tiny so there is a small number of metal cations and, subsequently, delocalised electrons.

10 A. High surface area to volume ratio means there is a large number of tiny metallic nanoparticles in a small volume.

Deconstructed questions

11 A **12** D

13 [The engineer should choose AISI 1010 carbon steel.[1]][In carbon steel, carbon atoms which are smaller than iron cations distort regular layers of the iron cations, meaning that it requires a lot of force to make iron cations slide over each other. Therefore, the fewer carbon atoms added, the less force required to deform carbon steel.[2]][AISI 1010 has the lowest percentage of carbon so it requires the least amount of force to bend and reshape this type of carbon steel. Hence, it is suitable for making the engineer's wires.[3]]

- ✓ ✗ I have identified the ideal carbon steel for the engineer to use.[1]
- ✓ ✗ I have explained how the percentage of carbon can affect the structure of carbon steel.[2]
- ✓ ✗ I have linked my answer to the question.[3]

Exam-style questions

With lesson

14 a [Stainless steel is a substitutional alloy[1]] [because iron and nickel atoms have relatively similar sizes.[2]]

- ✓ ✗ I have determined which type of alloy stainless steel is.[1]
- ✓ ✗ I have explained my answer.[2]

b [Even though iron and nickel atoms have similar sizes, there is still a small difference in their sizes. Therefore, nickel atoms slightly distort regular layers of iron cations in the alloy lattice, meaning that it requires more force to make all these metal cations slide over each other in the alloy lattice compared to pure iron lattice.[1]] [Moreover, nickel is also a metal so its cations are strongly attracted to delocalised electrons in the alloy lattice. Hence, the electrostatic force is stronger in the alloy lattice than in the pure iron lattice.[2]]

- ✓ ✗ I have explained the effect of atom size on the strength of stainless steel and why stainless steel is harder and stronger than pure iron.[1]
- ✓ ✗ I have described the effect of bonding on the strength of steel.[2]

15 a [In step 1, Steve should anneal the piece of iron.[1]] [Annealing can form large crystals which are flexible and easy to deform. Therefore, the softness and malleability of iron is increased and it's easier to shape it into a bicycle frame.[2]]

- ✓ ✗ I have identified the appropriate heat treatment.[1]
- ✓ ✗ I have explained how the heat treatment affects the malleability of the metal.[2]

b [In step 2, Steve should quench the bicycle frame.[1]] [Quenching forms small metallic crystals which are hard to break, and so the frame will be harder and stronger, and therefore able to withstand a heavy load.[2]]

- ✓ ✗ I have identified the appropriate heat treatment.[1]
- ✓ ✗ I have explained how the heat treatment affects the strength of the metal.[2]

16 a [All three types of metallic nanomaterials have nanoscale structures.[1]] [Metallic nanoparticles are spherical particles. Metallic nanorods are rod-like shaped structures. Metallic nanowires are wire-shaped structures.[2]]

- ✓ ✗ I have identified the similarity in the structures of the three types of metallic nanomaterials.[1]
- ✓ ✗ I have described the differences in the structures of the three types of metallic nanomaterials.[2]

b [A high heat conductivity allows metallic nanowires to be used to design energy-efficient buildings.[1]] [They can effectively absorb, store and release heat and therefore can replace air-conditioning systems that consume a large amount of electricity.[2]]

- ✓ ✗ I have identified the property that allows metallic nanowires to be used to design energy-efficient buildings.[1]
- ✓ ✗ I have described the property that allows metallic nanowires to be used to design energy-efficient buildings.[2]

17 [Type B metallic nanoparticles are the best for drug delivery[1]] [because they have a high surface area to volume ratio which enables them to help drug particles move through the bloodstream at a faster rate while maintaining a normal body temperature (37°C) which can protect healthy cells from being damaged.[2]] [Type A metallic nanoparticles have a lower surface area to volume ratio than type B metallic nanoparticles so they diffuse drugs at a lower rate. Type C metallic nanoparticles have a slightly higher surface area to volume ratio than type B metallic nanoparticles but they require a high temperature to be active, which can be dangerous for healthy cells.[3]]

- ✓ ✗ I have identified the best type of metallic nanoparticles for drug delivery.[1]
- ✓ ✗ I have described the property of the nanoparticle that is best suited to drug delivery in the body.[2]
- ✓ ✗ I have compared the nanoparticles against the other available options.[3]

Multiple lessons

18 a $m(\text{Fe}) = m(\text{steel}) \times \frac{90}{100} = 175 \times \frac{90}{100} = 157.5\ \text{g}$

$n(\text{Fe}) = \frac{157.5}{55.8} = 2.82 = 2.8\ \text{mol}$

b $m(\text{C}) = m(\text{steel}) - m(\text{Fe}) = 175 - 157.5 = 17.5\ \text{g}$

$n(\text{C}) = \frac{17.5}{12.0} = 1.46 = 1.5\ \text{mol}$

19 a Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ ([Ar] $4s^2 3d^6$ is also correct)

Ni: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$ ([Ar] $4s^2 3d^8$ is also correct)

b [An iron atom has a larger radius than a nickel atom.[1]] [Iron and nickel are in the same period in the periodic table. Going from the left to the right across a period in the periodic table, the positive nuclear charge increases so valence electrons are attracted more strongly towards the nucleus and therefore the radii of the atoms decrease.[2]] [As nickel has a greater nuclear charge, it would be able to pull the valence electrons closer to the nucleus, resulting in a smaller atomic radius than iron.[3]]

- ✓ ✗ I have identified the atom that has a larger radius.[1]
- ✓ ✗ I have explained my answer with reference to the radius trend in the periodic table.[2]
- ✓ ✗ I have linked my answer to the question.[3]

c [A nickel atom is more electronegative than an iron atom.[1]] [Iron and nickel are in the same period in the periodic table. Going from the left to the right across a period in the periodic table, the positive nuclear charge increases so the ability to attract electrons also increases, and therefore the electronegativity increases.[2]] [As a result, nickel is able to attract electrons more strongly towards itself compared to iron, making nickel more electronegative than iron.[3]]

- ✓ ✗ I have identified the atom that is more electronegative.[1]
- ✓ ✗ I have explained my answer with reference to the negativity trend in the periodic table.[2]
- ✓ ✗ I have linked my answer to the question.[3]

Key science skills

20 **a** The independent variable is the amount of the time the chocolate in the trays is left to cool down.

b The constant variables are the amount of chocolate in each tray, the type of chocolate, the length of time the chocolate is left on the pot and the size of the trays.

c [The chocolate in tray B had more time to cool than the chocolate in tray A. Therefore, we would expect the chocolate in tray B to snap very easily and cleanly in comparison to the chocolate in tray A.[1]] [However, we didn't get the expected result. This might be due to the fact that Warren didn't spread the chocolate in tray B evenly to the same extent as tray A and/or that he added more chocolate in tray B, which would then need more time to cool down.[2]]

✓ ✗ I have described the expected results.[1]

✓ ✗ I have given possible reasons for the obtained results.[2]

3C Reactivity of metals

Theory review questions

1 A. Not all but the majority of metallic elements in the periodic table can react with acid, water or oxygen.

2 A. Metals have low ionisation energies so electrons are easily taken from them.

3 B. Across the groups, ionisation energies of metals increases and therefore their reactivity with acid decreases.

4 B. Since a positive metal ion bonds to a negative ion from an acid to form a salt, free hydrogen ions in the solution bond to each other to produce hydrogen gas.

5 B. Metals with lower ionisation energies react faster with water, while metals with higher ionisation energies react slower with water.

6 A. A 'pop' sound is heard when hydrogen gas produced readily reacts with oxygen in the atmosphere in the presence of a flame.

7 A. Reaction between a metal and water will produce a metal hydroxide.

8 B. Rust is a result of a reaction between a metal and oxygen.

9 B. Some metals like gold and platinum cannot react with oxygen under normal conditions.

10 A. Zinc can react with oxygen to produce zinc oxides under normal conditions, while silver cannot.

Deconstructed questions

11 A **12** B

13 [Due to the humid condition in the laboratory, there was water in the room atmosphere.[1]] [Group 1 metals are highly reactive due to their low ionisation energies.[2]] [The small explosion was the result of the reaction between Group 1 metals with water in the atmosphere at room temperature with hydrogen gas produced.[3]]

✓ ✗ I have described the effect of the humidity on the room atmosphere.[1]

✓ ✗ I have explained the reactivity of Group 1 metals.[2]

✓ ✗ I have explained the underlying reason for the explosion observed.[3]

Exam-style questions

Within lesson

14 C

15 [Box A contains lead and box B contains calcium.[1]] [Lead has higher ionisation energy than calcium, meaning that it is more difficult for lead to lose electrons. Therefore, lead reacts slower with water and hydrogen gas bubbles were produced slower. Calcium has lower ionisation energy than lead, meaning that it is easier for calcium to lose electrons. Hence, calcium reacts more vigorously with dilute sulphuric acid and more bubbles were produced at a faster rate.[2]]

✓ ✗ I have determined the metal contained in each box.[1]

✓ ✗ I have explained the observation of the reaction between the metals and the acid.[2]

16 **a** [The student's method is incorrect.[1]] [Both potassium and zinc can react with sulfuric acid to produce colourless saline solution of salts and hydrogen gas.[2]]

✓ ✗ I have identified the accuracy of the student's method.[1]

✓ ✗ I have described the reaction of the metals with sulfuric acid.[2]

b [Potassium and zinc can be distinguished by placing the same amount of each metal into water at room temperature.[1]] [Potassium can react rapidly with water at room temperature to produce hydrogen gas, while zinc cannot react with water.[2]]

$2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$

$Zn_{(s)} + H_2O_{(l)} \rightarrow$ no reaction

✓ ✗ I have suggested one method to distinguish the two metals.[1]

✓ ✗ I have described the outcome of the method chosen.[2]

✓ ✗ I have included the possible balanced chemical equation(s) to support my answer.

17 [The engineer should choose iron.[1]] [Aluminium reacts more strongly with cold water than iron so using iron will help maintain better quality of agrimotors during the rain season.[2]]

✓ ✗ I have determined which metal should be chosen.[1]

✓ ✗ I have explained my answer with respect to the reactivity metals.[2]

Multiple lessons

18 **a** $1Cu_{(s)} + 2HNO_{3(aq)} \rightarrow 1Cu(NO_3)_{2(aq)} + 1H_{2(g)}$

b $n(Cu) = \frac{5.6}{63.5} = 0.0882$ mol

$n(H_2) = n(Cu) = 0.0882$ mol

$m(H_2) = 0.0882 \times 2.0 = 0.18$ g

19 **a** $2Fe_{(s)} + O_{2(g)} \rightarrow 2FeO_{(s)}$ (1)

$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$ (2)

b **i** $n(Fe) = \frac{4.2}{55.8} = 0.07527$ mol

$n(FeO) = n(Fe)$

$m(FeO) = 0.07527 \times (55.8 + 16) = 5.40$ g

ii $n(Fe) = \frac{7.8}{55.8} = 0.1398$ mol

$n(Fe_2O_3) = \frac{1}{2} \times n(Fe) = \frac{1}{2} \times 0.1398 = 0.0699$ mol

$m(Fe_2O_3) = 0.0699 \times (55.8 \times 2 + 16 \times 3) = 11.2$ g

Key science skills

20 a [The remaining solid mixture contains $Mg_{(s)}$ and $Cu_{(s)}$.[1]][This is because potassium can react rapidly with cold water to produce a hydroxide solution, whereas magnesium and copper cannot and therefore remain in the mixture.[2]]

$2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$

- I have determined the metals remained in the solid mixture.[1]
- I have justified my answer by comparing the ability to react with cold water of the metals.[2]
- I have included the occurred chemical reaction.

b [The student's suggestion is not correct.[1]][If concentrated sulfuric acid is placed into the mixture first, magnesium, potassium and copper will all react with the acid to produce a saline solution, meaning that we cannot collect pure solid magnesium.[2]]

- I have determined the accuracy of the student's suggestion.[1]
- I have explained my answer describing what would occur.[2]

c [The students collected less solid magnesium than expected because a small amount of magnesium slowly reacted with cold water in step 1 to produce magnesium hydroxide that stayed in the solution instead of the solid mixture.[1]][To collect more magnesium, they can decrease the waiting time after placing cold water into the mixture in step 1 to minimise the amount of magnesium reacting with cold water[2]][since potassium reacts really fast with cold water.[3]]

- I have explained the reason the students collected less magnesium than expected.[1]
- I have suggested one change that could be made to collect more magnesium.[2]
- I have described a reason to support my suggestion.[3]

Chapter 3 review

Multiple choice questions

1 D. The reactivity of metals depends on their ionisation energy.

2 D. The reactivity of metals increases going down a Group.

3 B. Rust forms when metals are exposed to oxygen over a certain period of time.

4 C. A high melting point is an indicator of the strength of the metallic bonds between metal atoms.

5 A. Substitutional alloys consist of metal atoms of a similar atomic radius.

6 A. Metals are able to conduct electricity due to the movement of delocalised electrons.

7 B. When metals and an acid react, hydrogen gas is produced.

8 B. The annealing of metals forms large metal crystals.

9 C. Reactivity depends on the ionisation energy of a metal.

10 C. The delocalised electrons in metals are responsible for many of their unique properties including lustre and heat conductivity.

Short answer questions

11 a [An ore is a deposit of minerals which may contain metal atoms such as iron, zinc and gold.[1]][Ores have to be purified to remove any sand and other debris first before the target metal is able to be extracted from pure samples of the ore.[2]]

- I have defined an ore.[1]
- I have described the process of extracting a metal from its ore.[2]

b [In a blast furnace, limestone and coke are added to the iron ore.[1]] [The hot air reacts with coke to produce carbon dioxide, and the limestone decomposes into calcium oxide and carbon dioxide.[2]] [The carbon dioxide produced then reacts with coke to produce carbon monoxide which reduces the iron ore, $Fe_2O_{3(s)}$, to form iron.[3]] [The iron produced sinks to the bottom and is extracted as a liquid while calcium oxide reacts with sand to form slag, which floats to the surface and is drained off.[4]]

- I have described the starting substances and equipment required of iron ore extraction.[1]
- I have described the production of carbon dioxide.[2]
- I have described the reduction of iron ore by carbon monoxide.[3]
- I have described the extraction of iron.[4]

c [A large-scale environmental impact of iron ore extraction is the production of carbon dioxide which contributes to global warming and the Greenhouse effect.[1]][A local environmental impact of iron ore extraction is the destruction of the land which may be of cultural significance by the digging of a giant mine in the ground.[2]]

- I have described an example of a large-scale environmental impact.[1]
- I have described an example of a local environmental impact.[2]

d [Some locals may argue that the mine and extraction centre will support the local economy by providing jobs and a source of income for the community.[1]][Other locals may argue that the mine could release harmful chemicals which have adverse effects on the local population's health.[2]]

- I have described an example of a positive societal effect of mining.[1]
- I have described an example of a negative societal effect of mining.[2]

e [Metals are malleable and strong, which allows them to be bent into different shapes which make them suitable for weight-bearing structures such as bridges.[1]][Metals are malleable because the metal cations can slide over each other and still be held together by strong electrostatic forces of attraction with the delocalised electrons.[2]] [Metals are strong due to the electrostatic force of attraction between delocalised electrons and metallic ions in the lattice.[3]]

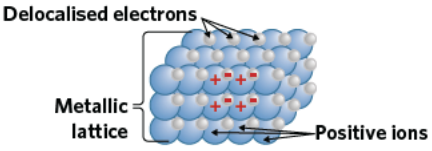
- I have identified the strength and malleability of metals.[1]
- I have explained why metals are malleable.[2]
- I have described the structure of metallic bonding that contributes to the strength of metals.[3]

12 a An alloy is a composite material consisting of a metal combined with another element which may be a metal or a non-metal.

b [One type of alloy is an interstitial alloy which consists of the alloying element spaced in between the principal metal cations.[1]][The other type of alloy is a substitutional alloy where some principal metal cations are replaced by the alloying element of similar atomic radius.[2]]

- I have identified and defined interstitial alloys.[1]
- I have identified and defined substitutional alloys.[2]

c [Metallic bonding consists of metal cations in a lattice structure surrounded by a 'sea' of delocalised electrons as shown in the diagram.[1]][In order to conduct electricity, there needs to be the flow of charged particles within the substance.[2]][In the case of metals, they can conduct electricity because of the 'sea' of delocalised electrons, which are able to flow throughout the metallic lattice, carrying charge and, therefore, electricity.[3]]



- I have described the model of metallic bonding.[1]
- I have described the requirements for a substance to conduct electricity.[2]
- I have justified why metals are able to conduct electricity.[3]
- I have included a diagram.

d [The wires of a pacemaker need to be malleable to be inserted through the irregularly shaped veins and arteries of the body.[1]] [Wires need to be able to conduct electricity to send the electrical pulses required for the heart to beat.[2]]

- I have identified a key reason for the need of the wires to be malleable.[1]
- I have identified a key reason for the need of wires to be able to conduct electricity.[2]

e [One of the key requirements of the wires in the pacemaker is that they need to be ductile.[1]][Annealing of the metal is a technique that improves the ductility and malleability of the metal, whereas quenching increases the hardness and tempering increases the brittleness – both properties which are not suitable to fulfil the wire's function.[2]] [Annealing is the process of heating a metal to its recrystallisation point and then cooling it down slowly, allowing large crystals to form.[3]]

- I have identified the key requirements needed for the wires in a pacemaker.[1]
- I have compared annealing to quenching and tempering.[2]
- I have explained the process of annealing with reference to recrystallisation.[3]

13 a [Barium would have the greatest reactivity.[1]][Although all three elements are found in the same group, barium contains more energy shells than beryllium and magnesium. As a result, the valence electrons are further away from the nucleus of the barium atom and are therefore held less strongly to the atom. Subsequently, valence electrons of barium are lost more easily, making it highly reactive.[2]]

- I have identified the element with the greatest reactivity.[1]
- I have explained the difference in reactivity based on the number of energy shells of the atom.[2]

b $Ba_{(s)} + 2H_2O_{(l)} \rightarrow Ba(OH)_{2(aq)} + H_{2(g)}$

[If a flame were placed inside the test tube, a 'pop' would be heard due to the ignition of hydrogen gas.[1]]

- I have included a balanced equation.
- I have identified the effect of placing a flame close to the reaction.[1]

c [Over time, metals exposed to the atmosphere will start to react with oxygen and start to form metal oxides and, therefore, the metals will start to crumble.[1]][This could be used to determine the reactivity of metals by comparing the times at which they react and form metal oxides.[2]][Barium is predicted to react first with oxygen according to the equation: $2Ba_{(s)} + O_{2(g)} \rightarrow 2BaO_{(s)}$.[3]][Magnesium will be predicted to react next with oxygen according to the equation: $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$.[4]][Finally, beryllium will be the last to react with oxygen according to the equation: $2Be_{(s)} + O_{2(g)} \rightarrow 2BeO_{(s)}$.[5]]

- I have identified what will happen when metals are left outside.[1]
- I have explained how this could be used to determine the reactivity of metals.[2]
- I have given the reaction between barium and oxygen.[3]
- I have given the reaction between magnesium and oxygen.[4]
- I have given the reaction between beryllium and oxygen.[5]

Key science skills questions

14 a [Caesium has to be kept in a vacuum due to its extreme reactivity with oxygen, as well as with water.[1]][Caesium reacts violently with oxygen according to the equation: $4Cs_{(l)} + O_{2(g)} \rightarrow 2Cs_2O_{(s)}$.[2]]

✓ ✗ I have explained why caesium has to be kept in a vacuum.[1]

✓ ✗ I have justified my answer with reference to its violent reaction with oxygen.[2]

b [Validity means whether the experimental technique can support the aim or, in this case, whether practicing with sodium is helpful before handling caesium.[1]][This is a valid technique because sodium is another highly reactive metal in Group 1, but it is less reactive than caesium and therefore is safer.[2]]

✓ ✗ I have defined validity.[1]

✓ ✗ I have identified whether this technique is valid.[2]

c [Accuracy relates to how close a value is to the true value, whereas precision relates to how close results are to each other.[1]][In this case, atomic clocks could be accurate but they don't have to be precise because different clocks could be accurately timekeeping but not showing the exact same reading (i.e. they may be in different locations).[2]]

✓ ✗ I have defined accuracy and precision.[1]

✓ ✗ I have explained why accuracy does not necessarily result in precision.[2]

4A Ionic compounds

Theory review questions

1 B. The charge of an ion and its magnitude are written as superscript after the chemical symbol.

2 B. Electrostatic force is formed between charged particles while atoms are neutral.

3 B. An ionic bond is formed only when the difference in electronegativity between a metal and a non-metal is equal or larger than two.

4 A. A single positive ion can attract multiple negative ions and a single negative ion can attract multiple positive ions.

5 A. Bohr models can show how ionic bonds are formed and the ratio of cations and anions but cannot demonstrate how ions are arranged in space.

6 B. In a sample of sodium chloride, each metal ion is bonded to adjacent negative non-metal ions if they are close enough.

7 B. Ionic bonds can be formed between a single cation and multiple neighbouring anions and vice versa, which forms a 3D structure of a sample of an ionic compound.

8 A. In a sample of sodium chloride, each sodium ion is bonded to six chloride ions and each chloride ion is bonded to six sodium ions.

9 B. The bond between each sodium ion and its six neighbouring chlorine atoms are all ionic bonds.

10 A. Cooling down more slowly means the crystals have more time to grow bigger.

Deconstructed questions

11 C **12** D

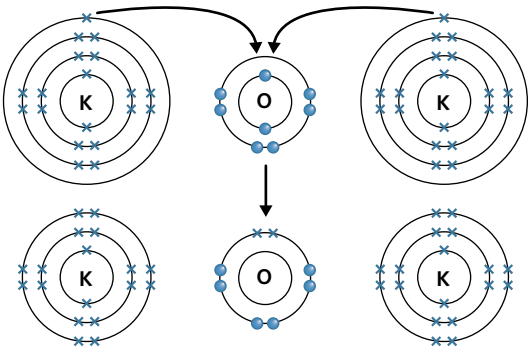
13 [The student's demonstration is incorrect[1]][because one F^- ion can be bonded to multiple Ca^{2+} ions and one Ca^{2+} ion can be bonded to multiple F^- ions, which was not shown in the student's model.[2]] [A three-dimensional lattice structure best represents ionic bonds present in CaF_2.[3]]

- ✓ ✗ I have determined the accuracy of the student's demonstration.[1]
- ✓ ✗ I have justified my answer comparing the student's demonstration with the nature of ionic bonds.[2]
- ✓ ✗ I have suggested a model that best represents ionic bonds.[3]

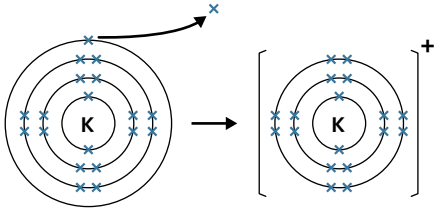
Exam-style questions

Within lesson

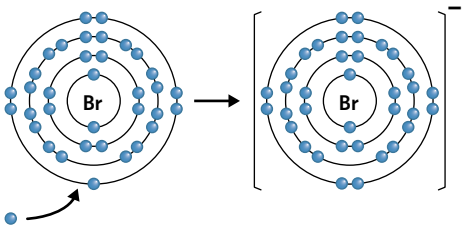
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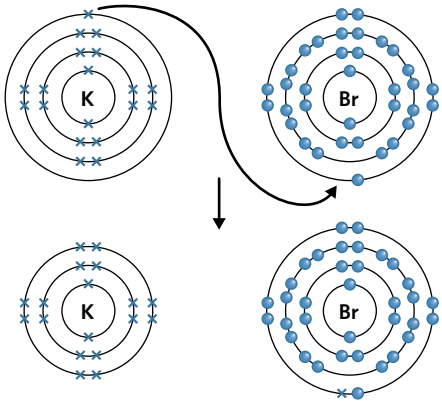
15 a



b



c



16 a [An ionic bond can be formed between calcium and chlorine[1]] [because the difference in electronegativity between calcium and chlorine is 2.16 which is large enough such that chlorine is able to attract electron(s) donated from calcium atoms.[2]] [This results in the formation of chloride anions and calcium cations that are attracted to form an ionic bond.[3]]

- ✓ ✗ I have determined if an ionic bond can be formed.[1]
- ✓ ✗ I have explained my answer using the difference in electronegativity.[2]
- ✓ ✗ I have linked my answer to the question.[3]

b [An ionic bond cannot be formed between carbon and chlorine[1]] [because the difference in electronegativity between carbon and chlorine is 0.61 which is not large enough for chlorine to attract electron(s) from carbon.[2]]

- ✓ ✗ I have determined if an ionic bond can be formed.[1]
- ✓ ✗ I have explained my answer using the difference in electronegativity.[2]

17 [The student's statement is incorrect.[1]] [The bond between a potassium ion and the oxide ion it donates its electrons to and the bonds between that potassium ion and the other three neighbouring oxygen ions are all ionic bonds. This is because ionic bonds can be formed between a single cation and multiple anions if they are close enough to each other.[2]]

- ✓ ✗ I have determined the accuracy of the student's statement.[1]
- ✓ ✗ I have justified my answer based on the condition of the formation of ionic bonds.[2]

Multiple lessons

18 **a** $Mg + 2HCl \rightarrow MgCl_2 + H_2$

b [Magnesium chloride is an ionic compound[1]][because it is formed by an ionic bond between a metallic cation (magnesium) and a non-metallic anion (chloride).[2]]

I have determined the type of compound.[1]

I have explained my answer using the elements present in the compound.[2]

c $n(Mg) = \frac{7.8}{24.3} = 0.321 \text{ mol} = 0.32 \text{ mol}$

19 **a** Metallic bonds are present in the sodium lattice. Ionic bonds are present in the sodium chloride lattice.

b [Ionic bonds are stronger than metallic bonds.[1]][As shown by the results, sodium chloride lattice has a higher melting point than sodium lattice, which means that it requires more energy to disrupt the ionic bonds present in sodium chloride lattice.[2]]

I have identified the stronger bond.[1]

I have explained my answer based on the difference in melting points.[2]

Key science skills

20 **a** Independent variable is the amount of salol added to each slide.

Dependent variable is the period of time of crystal formation.

b [Mikkel would see larger crystals on the heated slide.[1]][The heated slide requires more time to cool down and therefore more time for all crystals to be formed (80 seconds) while the cold slide requires less time for all crystals to be formed (20 seconds).[2]][As a result, larger crystals are able to be formed in the heated slide.[3]]

I have determined the slide with larger crystals.[1]

I have explained my answer using the duration in which crystals are formed in each slide.[2]

I have linked my answer to the question.[3]

4B Properties of ionic compounds

Theory review questions

1 A. The ionic bonds between cations and anions determine ionic compounds' properties.

2 B. Ionic bonds between cations and anions are strong and require a lot of energy to disrupt.

3 A. Ionic bonds between cations and anions are strong but when a strong force is applied, like-charged ions are pushed close to each other and the repulsive forces between them can break the ionic lattice.

4 B. Ionic bonds between cations and anions are strong so it requires a large amount of heat to disrupt them.

5 B. Ionic compounds can conduct electricity only when cations and anions move freely.

6 B. In aqueous state, cations and anions can move freely within the solution.

7 A. Many ionic compounds are very hard so they are used to make building stone in the construction industry.

8 B. A furnace can run with a very high temperature so a high melting point allows MgO to be the lining surface of a furnace.

9 B. KOH solution has K^+ cations and OH^- anions that move freely so it can conduct electricity.

10 B. Engines sometimes run with high temperature so high melting points allow some ionic compounds to be used to make engine components.

Deconstructed questions

11 B **12** C

13 [The students can dissolve the table salt in a cup of water and then connect the salt solution to the wires.[1]][Table salt is primarily made of sodium chloride which is an ionic compound. Therefore, it cannot conduct electricity in solid state due to the fact that all cations and anions are held in fixed positions via ionic bonds.[2]][However, when sodium chloride is dissolved in water, sodium cations and chloride anions can move freely within the solution and therefore allow for the electric charge to be carried through the circuit. As a result, sodium chloride in aqueous state can conduct electricity.[3]]

I have suggested one thing the students can do.[1]

I have explained my suggestion by explaining the compound's ability to conduct electricity in solid state.[2]

I have explained my suggestion by explaining the compound's ability to conduct electricity in aqueous state.[3]

Exam-style questions

Within lesson

14 [Limestone is made of $CaCO_3$ which is an ionic compound.[1]][It was not broken in the first three attempts because ionic bonds between Ca^{2+} ions and CO_3^{2-} ions are strong, meaning that the force applied was not strong enough to break the ionic bonds.[2]][However, when more force was applied, it's likely that the lattice became disrupted in a way that Ca^{2+} ions were pushed close to each other as well as the CO_3^{2-} ions. The repulsive forces between the like-charged ions disrupted the lattice structure of $CaCO_3$, causing the limestone to break.[3]]

I have identified the type of compound to which limestone belongs to.[1]

I have described the strength of ionic bonding.[2]

I have explained the effect of the disruption of the ionic lattice on the brittleness of limestone.[3]

15 [The student's explanation is incorrect.[1]][BaO has a high melting point due to strong ionic bonds between Ba^{2+} cations and O^{2-} anions, not as a result of the repulsive forces between similar ions.[2]]

I have identified the accuracy of the student's explanation.[1]

I have explained my answer with reference to the nature of ionic bonds.[2]

16 [Tap water contains freely moving ions.[1]][When a hair dryer is dropped into a bathtub filled with tap water, the ions found in water can move freely and thereby transfer the electric charge from the hair dryer to the human body. This results in an electrocution.[2]]

✓ ✗ I have identified the properties of tap water based on given data.[1]

✓ ✗ I have explained how water can cause electrocution based on the conduction of charge.[2]

Multiple lessons

17 [Sodium is a metallic substance and therefore contains delocalised electrons in its structure.[1]] [The delocalised electrons can move freely within the sodium lattice structure, resulting in the movement of electrons when the lattice is connected to an electrical circuit.[2]] [In contrast, in the sodium chloride lattice structure, ions are fixed in position and therefore are unable to move to carry electric charge.[3]] [Therefore, sodium lattice has a high value of electrical conductivity and sodium chloride lattice cannot conduct electricity.[4]]

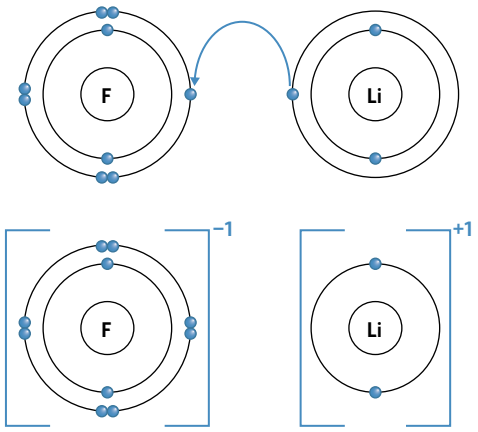
✓ ✗ I have identified the structure of sodium.[1]

✓ ✗ I have described the ability of sodium to conduct electricity.[2]

✓ ✗ I have compared the structure of sodium to that of sodium chloride and its ability to conduct electricity.[3]

✓ ✗ I have linked the difference in structure to the difference in electrical conductivity.[4]

18 a



b [Lithium fluoride has a high melting point due to the fact that the ionic bonds between Li^+ cations and F^- anions are strong.[1]] [Therefore, a large amount of heat energy is required to disrupt the lattice to melt lithium fluoride.[2]]

✓ ✗ I have described the ionic bonds between lithium and fluoride ions.[1]

✓ ✗ I have identified the effect of the bonds between lithium and fluoride ions on the melting point.[2]

19 a $n(Ca) = \frac{10.5}{40.1} = 0.262$ mol

b [Calcium chloride solution can conduct electricity because in aqueous state, Ca^{2+} cations and Cl^- anions can move freely, allowing for electric charge to be passed through the solution.[1]] [Solid calcium can also conduct electricity due to the delocalised electrons in its metallic lattice that can move freely to allow for the flow of an electric current.[2]]

✓ ✗ I have explained the electrical conductivity of the ionic compound in aqueous state.[1]

✓ ✗ I have explained the electrical conductivity of a metal in solid state.[2]

Key science skills

20 a The independent variable is the mass of solid NaCl used.
The dependent variable is the current recorded by the ammeter.

b [Solid NaCl cannot conduct electricity.[1]] [NaCl solutions in both beaker I and beaker II can conduct electricity. However, NaCl solution in beaker II has a higher electrical conductivity than the NaCl solution in beaker I.[2]]

✓ ✗ I have identified the electrical conductivity of an ionic compound in solid state.[1]

✓ ✗ I have identified the electrical conductivities of an ionic compound in aqueous state.[2]

c [NaCl solutions in both beaker I and beaker II can conduct electricity as both solutions contain Na^+ cations and Cl^- anions that are able to move freely in the solution.[1]] [NaCl solution in beaker II has a higher electrical conductivity than the NaCl solution in beaker I because more solid NaCl was dissolved into water in beaker II, meaning that beaker II has a greater number of Na^+ cations and Cl^- anions to carry the electric current.[2]]

✓ ✗ I have explained my conclusion using the movement of ions in aqueous solutions of an ionic compound.[1]

✓ ✗ I have explained my conclusion by comparing the number of moving ions in each case.[2]

4C Writing formulas for ionic compounds

Theory review questions

1 A. The cation is placed before the anion in the formula and name of an ionic compound.

2 B. Ionic compounds can form from any pair of cation and anion, no matter their magnitude of charge.

3 B. Lithium as pictured has lost one electron and is therefore positively charged.

4 B. The sulfate ion is an anion.

5 A. NH_4^+ is a polyatomic cation.

6 B. When naming an ionic compound, the cation (silver) precedes the anion (sulfate).

7 A. Brackets are not required around monatomic ions.

8 B. Double displacement reactions involve the transfer of ions between chemical species.

9 B. Balanced chemical equations have the same number of atoms of each element on either side of the equation.

10 A. The charge of ions with multiple different states possible is expressed through Roman numerals.

Deconstructed questions

11 A **12** D

13 $3Fe(SO_4)_2 + 2Fe(PO_4)_2 \rightarrow 2Fe(SO_4)_3 + Fe_3(PO_4)_4$

Exam-style questions

Within lesson

14

	Name	Formula
a	Silver	Ag^+
b	Copper (II)	Cu^{2+}
c	Hydroxide	OH^-
d	Fluoride	F^-
e	Nitrate	NO_3^-
f	Ammonium	NH_4^+
g	Sulfate	SO_4^{2-}
h	Iron (III)	Fe^{3+}
i	Phosphate	PO_4^{3-}
j	Sulfite	SO_3^{2-}

15

		Name	Formula
a	Na^+ and NO_3^-	Sodium nitrate	$NaNO_3$
b	Silver ion and sulfate ion	Silver sulfate	Ag_2SO_4
c	NH_4^+ and F^-	Ammonium fluoride	NH_4F
d	Hydrogen carbonate ion and strontium ion	Strontium hydrogen carbonate	$Sr(HCO_3)_2$
e	K^+ and O^{2-}	Potassium oxide	K_2O
f	Calcium ion and carbonate ion	Calcium carbonate	$CaCO_3$
g	Mg^{2+} and SO_3^{2-}	Magnesium sulfite	$MgSO_3$
h	Hydroxide ion and aluminium ion	Aluminium hydroxide	$Al(OH)_3$
i	Fe^{3+} and S^{2-}	Iron (III) sulfide	Fe_2S_3
j	Phosphate ion and Zn^{2+}	Zinc phosphate	$Zn_3(PO_4)_2$

16 **a** $Cu_2SO_4 + 2NaI \rightarrow 2CuI + Na_2SO_4$

b $NiO + FeSO_4 \rightarrow NiSO_4 + FeO$

c $FeN + 3NH_4HCO_3 \rightarrow Fe(HCO_3)_3 + (NH_4)_3N$

d $Ba_3(PO_4)_2 + 6HF \rightarrow 2H_3PO_4 + 3BaF_2$

e $Al_2(SO_3)_3 + 6AgNO_3 \rightarrow 2Al(NO_3)_3 + 3Ag_2SO_3$

Multiple lessons

17 **a** [$CaCO_3$.[1]][In terms of the structure of calcium carbonate, the ionic formula reveals that Ca^{2+} and CO_3^{2-} exist together in a ratio of 1 : 1 in calcium carbonate.[2]]

✓ ✗ I have identified the ionic formula.[1]

✓ ✗ I have described what the ionic formula reveals about the ratio of cations and anions.[2]

b [Ionic compounds are hard which makes them suitable as shells which function to protect marine organisms.[1]][Ionic compounds are hard because of the very strong electrostatic force of attraction between ions which makes them very hard to pull apart.[2]]

✓ ✗ I have identified the hardness of ionic compounds.[1]

✓ ✗ I have explained why ionic compounds are hard.[2]

c CaO

d [Ionic compounds decompose at very high temperatures.[1]] [This is because in order for an ionic compound to decompose, the strong electrostatic force of attraction between ions needs to be disrupted which takes a lot of energy.[2]]

✓ ✗ I have identified the high temperatures required to decompose ionic compounds.[1]

✓ ✗ I have explained the reason why hard temperatures are required.[2]

e [No.[1]][This is because in this equation ions are not exchanged between chemical species.[2]]

✓ ✗ I have answered the question.[1]

✓ ✗ I have justified my answer.[2]

18 **a** Potassium iodide

b NaCl

c $KI + NaCl \rightarrow KCl + NaI$

d $KI + AgNO_3 \rightarrow AgI + KNO_3$

This is a double displacement reaction.

e [Isotopes are atoms of the same element with a different number of neutrons and so ^{131}I is one isotope of iodine, but others exist with a different number of neutrons.[1]][The number of protons is the same across isotopes of the same element but isotopes have a different number of neutrons, and therefore different mass numbers.[2]]

✓ ✗ I have defined isotope.[1]

✓ ✗ I have identified similarities and differences amongst isotopes of the same element.[2]

f [Since radioactivity is caused by the release of a neutron from the nucleus, ^{131}I must be more unstable than the isotopes of iodine found in table salt and so will readily release a neutron from the nucleus, reverting to another isotope.[1]][Once a neutron has been ejected from the nucleus of the atom, the mass number decreases by 1 and so the mass number of radioactive isotopes decreases over time until they reach a stable form.[2]]

✓ ✗ I have explained the radioactivity of ^{131}I.[1]

✓ ✗ I have explained the effect of radioactivity on mass number.[2]

19 **a** [Fluorine atoms have an electron configuration of $1s^2 2s^2 2p^5$ and so are highly electronegative and reactive because they have a strong core charge and require one more electron to fill their valence shell.[1]] [Fluoride ions on the other hand have an electron configuration of $1s^2 2s^2 2p^6$ and so are not as reactive nor electronegative as they have a full valence shell of electrons.[2]]

✓ ✗ I have identified the electron configuration and properties of fluorine.[1]

✓ ✗ I have identified the electron configuration and properties of the fluoride ion.[2]

b Ca^{2+} – calcium ion

PO_4^{3-} – phosphate ion

F^- – fluoride ion

$Ca^{2+} : PO_4^{3-} : F^-$

5 : 3 : 1

c [Ionic compounds, of which fluorapatite is one, are hard which makes it suitable to function as tooth enamel in order to protect the sensitive nerve structures in a tooth.[1]][Ionic compounds are hard because of the strong electrostatic forces of attraction between ions which take a lot of energy to disrupt.[2]]

✓ ✗ I have identified the hardness of ionic compounds.[1]

✓ ✗ I have explained the hardness of ionic compounds.[2]

d M(fluorapatite) =

$5 \times 40.1 + 3 \times (31.0 + 4 \times 16.0) + 19.0 = 504.5 \text{ g mol}^{-1}$

e $n(Ca_5(PO_4)_3F) = \frac{0.25 \times 10^{-3}}{504.5} = 5.0 \times 10^{-7}$

$n(O) : n(Ca_5(PO_4)_3F) = 12 : 1$

$n(O) = 12 \times 5.0 \times 10^{-7} \text{ mol} = 5.9 \times 10^{-6} \text{ mol}$

$N(O) = 5.9 \times 10^{-6} \times 6.02 \times 10^{23} = 3.6 \times 10^{18}$ atoms

Key science skills

20 a Quantitative

b Independent variable: identity of element (X, Y or Z)

Dependent variable: charge of element (X, Y or Z)

c [This is a systematic error because it affects all results to the same degree.[1]][It could be rectified by determining the difference between true weight and recorded weight, and adding/subtracting this amount to all the data collected.[2]]

✓ ✗ I have identified the type of error.[1]

✓ ✗ I have suggested a way to correct this error.[2]

Chapter 4 review

Multiple choice questions

1 D. Ionic compounds are made up of cations and anions.

2 B. An ion can form ionic bonds with multiple oppositely charged ions.

3 A. Ions with like charges will be repelled, while ions with opposite charges will be attracted via electrostatic forces.

4 C. Letting ionic compounds cool for longer after heating will produce larger crystals in their structures.

5 C. Electrical conductivity requires the movement of charged particles, however ionic compounds as a solid are in a fixed lattice.

6 B. When like-charged ions are pushed closer together, the repulsive forces can be strong enough to shatter and break the lattice.

7 A. The high melting point of lithium fluoride enables it to be used as a component in engines.

8 C. Magnesium is the cation and is therefore written first, and magnesium ions have a charge of +2 while fluoride has a charge of −1, so two fluoride ions are needed.

9 D. Iron (III) has a charge of +3, and hydroxide ions are polyatomic ions with an overall charge of −1, so three hydroxide ions are required.

10 C. The charges of all ions must be balanced so that every product has a net charge of 0.

Short answer questions

11 a NaCl

b i $m(NaCl) = 0.13 \times (23.0 + 35.5) = 7.6$ g

7.6 > 6.0, therefore this is not a healthy amount of salt.

ii [The customer's hypothesis is likely incorrect.[1]][This is because ionic compounds are usually highly heat resistant and have a high melting point because of the strength of the ionic bonds.[2]][Therefore, exposing food to moderate heat would be insufficient to degrade the molecule, and it would therefore still have the same health effects.[3]]

✓ ✗ I have identified that the customer's hypothesis is likely incorrect.[1]

✓ ✗ I have explained that strong ionic bonds gives ionic compounds a high melting point.[2]

✓ ✗ I have connected my answer to the question.[3]

c [It would be expected that the salt block would shatter and break if sufficient force were applied.[1]][This is because ionic compounds are very hard and brittle.[2]][This is because if sufficient force were given such that like-charged ions were pushed closer to each other, the strength of the repulsive forces between these ions would cause the block to shatter.[3]]

✓ ✗ I have identified the effect of the application of force on the salt block.[1]

✓ ✗ I have identified the properties of ionic compounds.[2]

✓ ✗ I have explained the effect of a disruption of the lattice on the brittle nature of ionic compounds.[3]

d

Use	Ionic compound	Properties for this use
Lining surface of engines	Magnesium oxide (MgO)	High melting point
Human bones and teeth	Calcium phosphate ($Ca_3(PO_4)_2$)	Hardness
Electrolyte enabling current flow in cell batteries	Ammonium chloride (NH_4Cl)	Electrical conductivity

12 a [Reactions between ionic compounds are called double displacement reactions[1]][and they involve the exchange of ions between ionic compounds, such that each cation is bound to the anion of the other reactant ionic compound.[2]]

✓ ✗ I have identified the type of reaction between ionic compounds.[1]

✓ ✗ I have described the exchange of ions in double displacement reactions.[2]

b $Ba_3(PO_4)_2$, barium phosphate

c **i** $CaBr_2 + Cu_2SO_3 \rightarrow CaSO_3 + 2CuBr$

ii $Ag_2CO_3 + 2KF \rightarrow 2AgF + K_2CO_3$

iii $Cu_3N_2 + 3Na_2S \rightarrow 3CuS + 2Na_3N$

d **i** Calcium sulfite and copper (I) bromide

ii Silver fluoride and potassium carbonate

iii Copper (II) sulfide and sodium nitride

e **i** Metal

ii JF

iii Group 1

13 **a** [Calcium fluoride's formula is CaF_2.[1]][This suggests that in the lattice structure, there is a 1 : 2 ratio of calcium cations to fluoride anions.[2]]

✓ ✗ I have identified the formula for calcium fluoride.[1]

✓ ✗ I have described the structure of the calcium fluoride lattice with respect to its formula.[2]

b **i** [Stage 2 would have the greatest impact on the strength of the lattice.[1]][This is because at this stage, the temperature and length of time at which the molten ionic compound is allowed to cool will affect the size of the crystals formed in the final structure.[2]][This can affect the overall strength of the window.[3]]

✓ ✗ I have identified the relevant stage of the process.[1]

✓ ✗ I have described how this stage affects the structure of the product.[2]

✓ ✗ I have identified the overall effect on the final product.[3]

ii [The main difference between the two samples is the size of the crystals formed.[1]][Compared to sample 2, sample 1 has larger crystals, suggesting that it was given more time to cool down at a higher temperature,[2]][whereas sample 2 would have been cooled quickly at a lower temperature, resulting in smaller crystals.[3]]

✓ ✗ I have identified the main difference between the samples.[1]

✓ ✗ I have described the conditions resulting in sample 1.[2]

✓ ✗ I have compared the conditions of sample 1 to sample 2.[3]

c Sample 2

Key science skills questions

14 **a** Independent variable – the identity of the compound

Dependent variable – the interactions observed in the presence of a flame

b Controlled variables may include:

- Temperature of the flame
- Distance of sample from the flame
- Amount of time sample exposed to flame
- Amount of sample

c **i** [This is not a reliable experimental method.[1]][A reliable experiment is one that can reproduce the same results on separate occasions.[2]][As each student obtained different results, this method is not reliable.[3]]

✓ ✗ I have identified the experimental method as unreliable.[1]

✓ ✗ I have defined reliability.[2]

✓ ✗ I have connected the results to the question.[3]

ii Systematic error

REVIEW

5A Covalent bonding

Theory review questions

1 B. When atoms are chemically bonded together, they form molecules.

2 B. Covalent bonding relies on the sharing of electrons.

3 B. Each pair of shared electrons is equivalent to a single bond.

4 A. Electron dot formulas can be used to represent the valence electrons of both individual atoms and larger molecules.

5 B. Outer shell electrons are also referred to as valence electrons.

6 A. Electron dot formulas show electrons in the valence shell.

7 B. The ball and stick model does not depict the relative sizes and scales of atoms.

8 B. Some atoms, such as sulphur, do not obey the octet rule.

9 B. Since there are 6 atoms bonded to the central atom, there are 6 single bonds which means the valence shell has 12 electrons.

10 B. The space-filling model depicts the relative scale of atoms.

Deconstructed questions

11 A

12 A

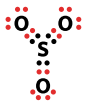
13 

Exam-style questions

Within lesson

14 a 

b 

c 

d 

e 

15 a Similarity: both representations show the valence electrons of each atom.

Difference: valence structure representation replaces bonding electrons with lines.

b Similarity: both models represent molecules in three-dimensions.

Difference: the space-filling model accurately depicts the size of atoms and distances between them.

c [Electron dot diagrams, valence structures and structural formulas are two-dimensional representations of molecules and so would be inappropriate for looking at the three-dimensional shape.[1]][Since the molecule has large atoms, if a space-filling model was used, a clear picture of the molecule would not be able to be generated due to all the different overlapping atoms.[2]][Therefore, the ball-and-stick model would be most appropriate.[3]]

- ✓ ✗ I have explained which models are not suitable for representing a molecule in three-dimensions.[1]
- ✓ ✗ I have explained the limitation of the space-filling model with respect to this application.[2]
- ✓ ✗ I have identified the best model.[3]

16 a 

b 

c [Oxygen has six valence electrons and so in order to obey the octet rule, must form two single bonds with its two unpaired electrons.[1]][Therefore, it usually has two lone pairs of electrons because four of its original electrons are not involved in bonding.[2]]

- ✓ ✗ I have justified my answer with reference to the octet rule.[1]
- ✓ ✗ I have linked the octet rule to the number of lone pairs of electrons and bonding electrons.[2]

17 a [Oxygen has six valence electrons and therefore, it wants another atom or atoms to share two electrons with it.[1]][In this case, since both oxygen atoms would like another atom to share two electrons with it, a double covalent bond would form as each oxygen atom shares two electrons with the other oxygen atom.[2]]

- ✓ ✗ I have identified the number of electrons needed for oxygen to satisfy the octet rule.[1]
- ✓ ✗ I have described the type of covalent bond that would form.[2]

b [Oxygen has six valence electrons and therefore, it wants another atom or atoms to share two electrons with it.[1]][In this case, since both oxygen atoms need to be bonded to one hydrogen atom and the other oxygen atom, two single covalent bonds will form (O-H and O-O). This means the oxygen-oxygen bond will be a single covalent bond.[2]]

- ✓ ✗ I have identified the number of electrons needed for oxygen to satisfy the octet rule.[1]
- ✓ ✗ I have described the type of covalent bond that would form.[2]

c [Double covalent bonds are stronger than single covalent bonds because there are more electrons shared in it.[1]][This means that the oxygen atoms are bonded to each other more strongly in O_2 than they are in H_2O_2.[2]][This would explain why oxygen gas (O_2) is more stable than hydrogen peroxide (H_2O_2), as shown by the fact that H_2O_2 decomposes explosively while O_2 is more stable in nature.[3]]

- ✓ ✗ I have explained the relative strengths of double and single covalent bonds.[1]
- ✓ ✗ I have compared the strength of the covalent bonds present in O_2 and H_2O_2.[2]
- ✓ ✗ I have compared the stability of O_2 and H_2O_2.[3]

18 a [Neon cannot form covalent bonds.[1]][This is because it has a full valence shell already and therefore, does not need to share any electrons with another atom to complete its octet.[2]]

- ✓ ✗ I have identified the number of covalent bonds neon can form.[1]
- ✓ ✗ I have justified my answer with reference to the octet rule.[2]

b [Oxygen has six valence electrons.[1]][Therefore, it would like another atom or atoms to share two electrons with it.[2]][As a result, this can either be in the form of one double covalent bond or two single covalent bonds.[3]]

Note: While not required for this response, carbon monoxide (CO) is an exception to the normal rule because the oxygen atom forms a triple covalent bond with the carbon atom.

✓ ✗ I have identified the number of valence electrons that oxygen has.[1]

✓ ✗ I have used the octet rule to justify the number of electrons it needs to complete its valence shell.[2]

✓ ✗ I have identified the different bonds that can be formed.[3]

c [Carbon has four valence electrons.[1]][Therefore, it would like another atom or atoms to share four electrons with it.[2]][As a result, this can either be in the form of one triple bond and one single bond, one double bond and two single bonds, two double bonds or four single bonds.[3]]

✓ ✗ I have identified the number of valence electrons that carbon has.[1]

✓ ✗ I have used the octet rule to justify the number of electrons it needs to complete its valence shell.[2]

✓ ✗ I have identified the different bonds that can be formed.[3]

d [Carbon is able to form a number of different types of covalent bonds with other atoms.[1]][Consequently, since it can form a number of different types of covalent bonds with other atoms, this will lead to an increase in the possible number of carbon-containing compounds that can be formed.[2]]

✓ ✗ I have referenced carbon's ability to form different types of covalent bonds.[1]

✓ ✗ I have linked this ability to the diversity of carbon-containing compounds.[2]

Multiple lessons

19 a [Covalent bonds are the chemical bond produced when one atom shares electron(s) with another atom[1]][whereas ionic bonds are the chemical bonds formed between a positively charged atom and a negatively charged atom, after a complete transfer of electrons rather than the sharing of electrons.[2]]

✓ ✗ I have defined a covalent bond.[1]

✓ ✗ I have defined an ionic bond with reference to a covalent bond.[2]

b [CCl_4 is covalently bonded and NaCl is ionically bonded.[1]][Covalent bonding is between non-metallic atoms (Na is a metal). Ionic bonding results when the difference in electronegativities is greater than 2 which is the case for Na and Cl.[2]]

✓ ✗ I have identified the type of bonding occurring within each compound.[1]

✓ ✗ I have used the characteristics of each bond type to explain my conclusion.[2]

c Determine whether or not the compound is able to dissolve in water and conduct electricity

Determine the melting point of each compound

Key science skills

20 a The independent variable is the identity of the chemical substance in the capillary tube.

b The aim of this experiment is to determine the type of bonding present by performing a melting point determination.

c Quantitative

d [Substance A is more likely to be a covalent compound than substance B.[1]][This is because its melting point, a measure of the force of attraction between molecules, is much lower than substance B. This is because covalent compounds generally have lower melting points compared to ionic compounds because it generally takes more energy to disrupt the electrostatic attraction in ionic compounds compared to the intermolecular forces between covalent compounds.[2]]

✓ ✗ I have identified which substance is more likely to be a covalent compound.[1]

✓ ✗ I have justified my choice by comparing the energy required to break the different bonds.[2]

5B Properties of covalent compounds

Theory review questions

1 A. The shape depicted has four atoms/pairs of electrons spaced equally around a central atom.

2 B. Lone pairs of electrons are not counted when determining the final shape of a molecule.

3 B. Double bonds, when it comes to determining the shape of a molecule, are considered as one pair of electrons.

4 A. Polar covalent bonds arise due to a large difference in electronegativity.

5 A. A non-polar covalent bond can be formed because the difference in electronegativity is negligible.

6 A. Not all compounds containing polar covalent bonds are considered polar.

7 B. VSEPR and bond polarity are both used to determine polarity.

8 A. Tetrahedral shapes are symmetrical.

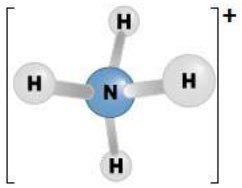
9 B. Symmetry affects the polarity of a molecule.

10 A. Symmetry affects the polarity of a molecule.

Deconstructed questions

11 C **12** B

13 [Both molecules have polar covalent bonds, N-H and O-H.[1]][H_2O has a bent (V-shaped) configuration which means it is not symmetrical in three-dimensions and the polarities don't cancel. This means it is polar.[2]][On the other hand, NH_4^+ has a tetrahedral shape. This means it is symmetrical in three-dimensions and so while it is charged (has a positive charge), it is non-polar.[3]][Since only non-polar compounds are able to enter the cell, NH_4^+ will be able to enter the cell but not H_2O.[4]]





- I have identified the polarity of the covalent bonds present in both molecules.[1]
- I have described the shape of the H_2O molecule and its overall polarity.[2]
- I have described the shape of the NH_4^+ molecule and its overall polarity.[3]
- I have justified which molecule will be able to enter the cell.[4]
- I have included diagrams in my answer.

Exam-style questions

Within lesson

14 B **15** D

16 [The electron clouds represent the density of electrons around each atom.[1]][When comparing molecules A and B, we can see that molecule B has an even density of electrons around both atoms whereas molecule A has an uneven electron density around each atom, where the atom found on the right is shown to have a greater density than the atom on the left.[2]][This suggests that the atom on the right of molecule A is more electronegative than the atom on the left, resulting in an uneven distribution of the shared pair of electrons.[3]][Considering that both molecules are considered as diatomic and therefore consist of only the two atoms shown, molecule A can be considered as being a polar molecule.[4]]

- I have identified the function of electron clouds.[1]
- I have described the electron distribution of both molecules.[2]
- I have described the electron distribution within molecule A.[3]
- I have linked my answer to the question.[4]

17 :C≡O:
Carbon monoxide

:Ö=C=Ö:
Carbon dioxide

[At a bond level, the covalent bonds between carbon and oxygen are polar covalent bonds due to the difference in electronegativity between carbon and oxygen.[1]][According to VSEPR, carbon monoxide is a linear molecule and is therefore, not symmetrical.[2]][As a result, this causes the molecule to have a partial negative end (located at the oxygen atom) and a partial positive end (located at the carbon atom), resulting in the molecule being polar.[3]][According to VSEPR, carbon dioxide is also a linear molecule.[4]][However, the symmetrical nature of the carbon dioxide molecule means that all polar bonds counterbalance each other, and therefore, there is no overall dipole moment experienced by the molecule. As a result, carbon dioxide is not polar.[5]]

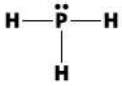
- I have identified the nature of the bond between carbon and oxygen.[1]
- I have described the structure of carbon monoxide.[2]
- I have explained the overall polarity of carbon monoxide.[3]
- I have described the structure of carbon dioxide.[4]
- I have explained the overall polarity of carbon dioxide.[5]
- I have included diagrams in my answer.

Multiple lessons

18 a



Electron dot diagram



Valence structure

b [Antony's conclusion is incorrect.[1]][This is because he has ignored the presence of the lone pair of electrons when determining the shape of the molecule.[2]][As seen in the electron dot diagram, the shape of the molecule is trigonal pyramidal.[3]][Since it is not symmetrical in three-dimensions, the molecule is polar (the lone pair of electrons also contributes to the molecule's polarity).[4]]

- I have identified the accuracy of Antony's statement.[1]
- I have identified a possible reason for the error in Antony's conclusion.[2]
- I have described the shape of phosphine.[3]
- I have linked my answer to the question.[4]

c [Due to the flammable nature of phosphine gas, it should be stored in a flame/ignition source free zone.[1]][Since phosphine is fatal when inhaled, it should be stored in leak proof containers with sensors for detecting leaks.[2]]

- I have recommended one way to improve safety due to the flammability of phosphine.[1]
- I have recommended one way to improve safety due to the fatal nature of phosphine when inhaled.[2]

19 a

Number of lone pairs on central atom	Number of single bonds	Ball and stick model	Shape
1	3		Trigonal pyramidal
2	2		Bent or V-shaped

b [A molecule with a trigonal planar shape is possible if the central atom does not obey the octet rule.[1]][In this case, there would be three pairs of electrons spaced 120 degrees around the central atom in a trigonal planar arrangement.[2]]

- I have identified an exception to the octet rule.[1]
- I have explained how a trigonal planar shape could arise.[2]

c [There is not just one combination of bonds and lone pairs of electrons that gives rise to a linear shape.[1]][This is because a molecule with one single bond and three lone pairs of electrons can form a linear shape, just like a molecule with two lone pairs of electrons and a double bond.[2]]

✓ ✗ I have identified ways in which a linear shape can be achieved.[1]

✓ ✗ I have justified my answer with reference to an example.[2]

d Electron pair repulsion.

e [Lone pairs of electrons are important in determining the shape of molecules in that they determine what arrangement the electrons take up.[1]][However, they are not included in the final shape.[2]]

✓ ✗ I have explained how lone pairs affect the shape of a molecule.[1]

✓ ✗ I have identified the effect of lone pairs on the final shape.[2]

Key science skills

20 a The shape of the molecule.

b [The results of this experiment are not valid.[1]][This is because validity is whether an experiment can prove the aim and in this case, since there are other confounding variables like the atoms in each molecule, the results are not valid.[2]]

✓ ✗ I have identified the validity of the results.[1]

✓ ✗ I have justified my answer with reference to the definition of validity.[2]

5C Intermolecular bonding

Theory review questions

1 B. Intermolecular bonding occurs between molecules, not atoms.

2 A. The number of electrons can affect an atom's ability to form instantaneous dipole moments, leading to stronger intermolecular bonding.

3 B. Dispersion forces occur due to electrostatic attraction between instantaneous dipoles.

4 A. As all atoms can form instantaneous dipoles, all molecules take part in dispersion forces.

5 B. The molecules presented are non-polar covalent compounds, which form dispersion forces.

6 A. Due to their permanent nature, dipole-dipole interactions are stronger than dispersion forces and only occur when molecules contain polar covalent bonds.

7 A. The greater the difference in electronegativity between bonded atoms, the more polar the compound. This results in a stronger dipole-dipole interaction. Compared to HCl, HF is more polar and therefore has stronger intermolecular forces.

8 B. In addition to dispersion forces, water molecules also have the ability to take part in hydrogen bonding, which requires more energy to break than dipole-dipole interactions and dispersion forces alone.

9 B. Intramolecular bonds are much stronger than intermolecular bonds and, therefore, would require more energy to break.

10 B. A larger surface area allows for more interactions between molecules.

Deconstructed questions

11 C **12** C

13 [As all of the molecules listed are non-polar covalent compounds, the only intermolecular bonds that exist between the molecules are dispersion forces.[1]][When comparing the molecules, the main difference between them is the size of the atoms present in the molecule. Larger molecules have more electrons, which results in stronger dispersion forces. Considering fluorine and chlorine exist as gases at room temperature, this suggests that these molecules have the weakest dispersion forces as the energy available is enough to surpass the boiling point of these molecules.[2]][When comparing bromine and iodine, iodine has a larger atomic mass and therefore has more electrons present in its structure. As a result, the dispersion forces holding iodine molecules together are stronger than bromine and so it exists as a solid at room temperature rather than as a liquid, like bromine. This suggests that the energy available is not sufficient to meet the melting point of iodine although it is enough to exceed the melting point of bromine.[3]]

✓ ✗ I have identified the similarities between the intermolecular bonds of all the molecules.[1]

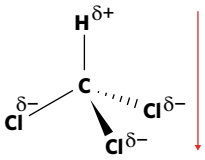
✓ ✗ I have related atomic size to the boiling points of the molecules.[2]

✓ ✗ I have related atomic mass and electron density to the melting points of the molecules.[3]

Exam-style questions

Within lesson

14 a



b [Dispersion forces are present in trichloromethane as they are found between all molecules.[1]][Dispersion forces arise from instantaneous dipole-dipole moments which are created from the constantly changing distribution of electrons.[2]][Permanent dipole-dipole forces are also present.[3]][This is because the molecule has an overall permanent dipole moment as shown in part a, which means that there is the electrostatic attraction between the positive and negative ends of $CHCl_3$ molecules.[4]]

✓ ✗ I have identified the presence of dispersion forces.[1]

✓ ✗ I have explained the origin of dispersion forces.[2]

✓ ✗ I have identified the presence of permanent dipole-dipole forces.[3]

✓ ✗ I have explained the origin of permanent dipole-dipole forces.[4]

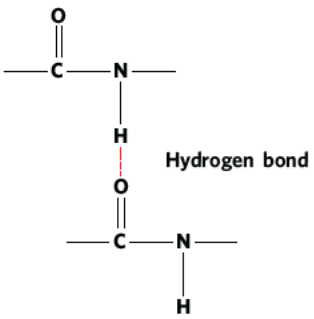
c [A substance's state of matter depends on the strength of intermolecular forces present.[1]][Since chloroform has both dispersion forces (a relatively weak intermolecular force) and permanent dipole-dipole forces (a stronger intermolecular force), there is sufficient intermolecular bonding for the molecules to be held together as a liquid at room temperature.[2]][However, there are insufficient forces of attraction for chloroform to exist as a solid at room temperature and the strong intermolecular forces present mean it can't exist as a gas at room temperature.[3]]

- ✓ ✗ I have linked a substance's state of matter to the strength of intermolecular forces present.[1]
- ✓ ✗ I have listed the intermolecular forces present in chloroform and have explained why chloroform is a liquid at room temperature.[2]
- ✓ ✗ I have justified why chloroform is neither a solid nor a gas at room temperature based on the strength of the intermolecular forces present.[3]

15 a [Hydrogen bonds are a much stronger form of intermolecular bonding than dispersion forces.[1]][Since protein stability is important, the strength of hydrogen bonds ensures that the protein can remain stable and not be disturbed as easily.[2]]

- ✓ ✗ I have compared the strength of hydrogen bonds to dispersion forces.[1]
- ✓ ✗ I have linked this comparison to the importance of hydrogen bonding in protein structures.[2]

b



c [An increase in temperature would increase the average kinetic energy of molecules, disrupting the hydrogen bonding present.[1]][Since hydrogen bonding helps to stabilise protein structure, a disruption in hydrogen bonding would affect structure and, therefore, disrupt the function of the protein.[2]]

- ✓ ✗ I have identified one way to disrupt intermolecular forces.[1]
- ✓ ✗ I have linked this effect to protein function.[2]

16 a Dispersion forces

b Dispersion forces arise from instantaneous dipole-dipole moments which are created from the constantly changing distribution of electrons.

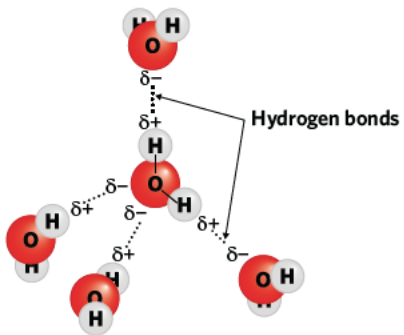
c [Based on the structure of molecules A and B, both have the same molecular formula and form non-polar covalent bonds.[1]][The main difference between the two molecules is the structure because molecule A is linear whereas molecule B is branched.[2]][When bonding to neighbouring molecules, molecule A has less branching and therefore forms stronger dispersion forces than molecule B as the molecules can pack more closely together.[3]][As a result, the melting point of molecule A would be higher than molecule B.[4]][Based on the results, molecule A most likely has a melting point of −130°C whereas molecule B most likely has a melting point of −160°C.[5]]

- ✓ ✗ I have identified the bonds in both molecules.[1]
- ✓ ✗ I have identified the key difference between both molecules.[2]
- ✓ ✗ I have described the effect of molecular structure on bond strength.[3]
- ✓ ✗ I have identified the relationship between melting point and branching.[4]
- ✓ ✗ I have linked my answer to the question.[5]

d [Boiling point is dependent on the strength of the intermolecular forces present.[1]][This is because the boiling point is where the intermolecular forces are disrupted to the point where the molecule transitions from a liquid to a gas.[2]][Hardness is also dependent on the strength of intermolecular forces present.[3]][This is because hardness is a measure of how well molecules are held together in a rigid structure.[4]]

- ✓ ✗ I have identified boiling point as being affected by intermolecular forces.[1]
- ✓ ✗ I have justified why boiling point is affected by intermolecular forces.[2]
- ✓ ✗ I have identified hardness as being affected by intermolecular forces.[3]
- ✓ ✗ I have justified why hardness is affected by intermolecular forces.[4]

17 a



- ✓ ✗ I have identified the type of bond occurring between the water molecules.
- ✓ ✗ I have labelled the polarity of each atom in the molecule.

b [When the temperature of water reaches its freezing point, the water molecules start to arrange themselves into a regular lattice.[1]] [This lattice structure means that the water molecules are spaced further apart from each other.[2]] [Since the water molecules are spaced further apart, ice becomes less dense than water, allowing it to float.[3]]

- I have identified the change in interaction between water molecules at freezing point.[1]
- I have explained this effect on the spacing of water molecules.[2]
- I have linked this decrease in density to ice being able to float.[3]

c [Water contains polar covalent bonds and, therefore, is a polar molecule, whereas the molecule given has only non-polar covalent bonds and therefore is a non-polar molecule.[1]] [As a result, water is able to take part in permanent dipole-dipole interactions, in particular hydrogen bonds, whereas the molecule shown can only take part in dispersion forces.[2]] [As hydrogen bonds are stronger than dispersion forces, water would have a higher melting point than the molecule shown since more energy is required to break these bonds.[3]]

- I have described the polarity of each molecule.[1]
- I have identified the type of intermolecular bonds for each molecule.[2]
- I have linked my answer to the question.[3]

d [Tetradecane is a very large molecule with 14 carbon atoms and 30 hydrogen atoms which means there are 114 electrons in total. As a result, tetradecane molecules form many intermolecular dispersion forces.[1]] [Tetradecane therefore has a higher melting point than water because the dispersion forces between tetradecane molecules are much stronger than the combined dispersion forces, permanent dipole-dipole bonds and hydrogen bonds between water molecules.[2]]

- I have described the size of tetradecane and its effect on the strength of the dispersion forces present.[1]
- I have explained why a higher melting point is possible.[2]

18 a Dispersion forces

b [As nitrogen gas, N_2, is non-polar, the molecules are held together by dispersion forces only.[1]] [Of the three main forces of intermolecular attraction, dispersion forces are the weakest and, therefore, require the least amount of energy to break.[2]] [As seen in the description, the boiling point of liquid nitrogen is very low, only −195.8°C, which illustrates the weakness of the dispersion forces present.[3]]

- I have identified the type of intermolecular forces.[1]
- I have compared the relative strength of intermolecular forces.[2]
- I have used evidence to support my conclusion.[3]

c [The average human body temperature is approximately 37°C. As such, when Angela touches the liquid nitrogen, the temperature of her finger is a lot higher than the boiling point of liquid nitrogen.[1]] [Therefore, the liquid nitrogen would turn into gas and escape from her hands.[2]] [Due to this, Angela would not experience any injury as the liquid nitrogen would not be in contact with her finger for long enough to cause harm.[3]]

- I have described the difference in temperature between Angela's hand and the boiling point of liquid nitrogen.[1]
- I have described the effect of Angela's temperature to the sample.[2]
- I have linked my answer to the question.[3]

Multiple lessons

19 a [The ionic bonding in a solid crystal of NaCl is an intramolecular force.[1]] [This is because ionic bonds exist between Na^+ and Cl^- within a regular lattice which is equivalent to one molecule.[2]]

- I have identified the ionic bonding in a solid crystal of NaCl as intramolecular.[1]
- I have explained my answer with reference to the definition of intramolecular bonding.[2]

b [Water is a polar compound.[1]] [Therefore, when ionic substances are placed into water, the oppositely charged poles of water are attracted to the positively charged sodium ions or the negatively charged chloride ions.[2]] [This would cause the sodium chloride to dissolve in water, resulting in a solution with charged particles.[3]]

- I have identified the polarity of water.[1]
- I have described the interaction between water and sodium chloride.[2]
- I have identified the effect of the interaction between water and sodium chloride.[3]

c [In solid form, sodium ions and chloride ions are held tightly together in a crystal lattice and therefore electricity has no ability to flow.[1]] [In a solution however, both sodium and chloride exist in free-flowing ion form.[2]] [The ability of these charged particles to move allows electricity to be conducted through a sodium chloride solution but not a sample of solid sodium chloride.[3]]

- I have described the nature of solid sodium chloride.[1]
- I have described the nature of sodium chloride in solution.[2]
- I have linked my answer to the question.[3]

d [HCl is a polar covalent molecule which means its intermolecular forces consist of dispersion forces and permanent dipole-dipole bonds.[1]] [NaCl is an ionic compound and ionic compounds are held together by strong, electrostatic forces of attraction between cations and anions.[2]] [Since electrostatic forces of attraction are much stronger than dispersion forces and permanent dipole-dipole bonds, NaCl will have a higher melting point as its molecules are held more strongly together.[3]]

I have described the intermolecular forces present in HCl.[1]

I have described the intramolecular forces present in NaCl.[2]

I have identified which compound has the highest melting point according to the intermolecular forces present.[3]

e i $Mg_{(s)} + Cl_{2(g)} \rightarrow MgCl_{2(s)}$

ii $n(Mg) : n(MgCl_2) = 1 : 1$

$n(MgCl_2) = \frac{1}{1} \times 2.0 = 2.0$ mol

Key science skills

20 a The substance used.

b To determine the effect of the structure of a molecule on its boiling point.

c [The results suggest that substance A has a higher boiling point than substance B.[1]][The fact that two different masses of substance were used (5 g of substance A and 4 g of substance B) would have decreased the validity of the results and lead to an incorrect conclusion being drawn.[2]][This is due to the fact that there was more than one variable that was changed during the experiment (multiple independent variables).[3]]

I have identified the conclusion of the experiment.[1]

I have identified the key error affecting the validity of the experiment.[2]

I have described the effect of uncontrolled variables on the validity of the experiment.[3]

Chapter 5 review

Multiple choice questions

1 C. Covalent bonds occur between non-metallic atoms.

2 D. The more electrons that are shared in a covalent bond, the stronger the force of attraction between the atoms and the greater the bond strength.

3 B. The space-filling model depicts the actual size and shape of atoms in a molecule.

4 D. The only difference between electron dot diagrams and valence structures is that in valence structures, bonding pairs of electrons are replaced by lines.

5 C. C_2H_2 has a triple covalent bond meaning that both carbon atoms have a linear arrangement of electron pairs.

6 D. Molecules with a permanent dipole are polar which means they have one or more polar covalent bonds present and are asymmetrical.

7 D. Molecule D is largely non-polar resulting in almost all of its intermolecular forces being dispersion forces.

8 C. Some molecules with polar covalent bonds are non-polar.

9 D. Hydrogen bonds occur when a hydrogen atom is 'sandwiched' between two electronegative atoms (either fluorine, oxygen or nitrogen).

10 B. The strength of intermolecular forces contributes to the boiling point of compounds.

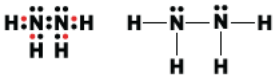
Short answer questions

11 a 

b [According to the chemical equation, one molecule of $N_{2(g)}$ reacts with two molecules of $H_{2(g)}$ to produce one molecule of $N_2H_{4(l)}$.[1]] [Therefore, to produce two molecules of $N_2H_{4(l)}$, two molecules of $N_{2(g)}$ and four molecules of $H_{2(g)}$ are required.[2]]

I have referenced the chemical equation.[1]

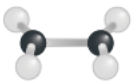
I have identified the ratio of products to reactants to calculate the answer.[2]

c 

I have drawn the electron dot diagram of hydrazine.

I have drawn the valence structure of hydrazine.

d [Each nitrogen atom has one lone pair of electrons, and three pairs of bonding electrons.[1]][As a result of this combination, according to VSEPR, each nitrogen atom will have a trigonal pyramidal shape.[2]] [Hence, the overall shape of the molecule is trigonal pyramidal as shown:[3]]



I have identified the number of lone pairs and bonding electrons of each nitrogen atom.[1]

I have identified the shape of each nitrogen atom according to VSEPR.[2]

I have described the three-dimensional shape of hydrazine.[3]

I have included a possible three-dimensional shape of hydrazine.

e [Hydrazine has polar N-H covalent bonds present.[1]][Since hydrazine is not symmetrical, these permanent dipole moments do not cancel out and therefore, it is a polar molecule.[2]]

I have identified the polar covalent bonds present in hydrazine.[1]

I have determined the polarity of the molecule.[2]

f [The molecule with the greatest boiling point would have the greatest strength of intermolecular forces.[1]][Since C_2H_4 is non-polar, there will be only dispersion forces present whereas N_2H_4 is polar with dispersion forces, permanent dipole-dipole forces and hydrogen bonds.[2]][Dipole-dipole forces and hydrogen bonds are stronger than dispersion forces and therefore would require more energy to break, resulting in N_2H_4 having a higher boiling point.[3]]

- I have identified the relationship between boiling point and intermolecular forces.[1]
- I have identified the intermolecular forces present in C_2H_4 and N_2H_4.[2]
- I have described the effect of intermolecular forces on the boiling point and have identified which molecule has the highest boiling point.[3]

12 a H—H :Ö=Ö: :N≡N:

- I have drawn the valence structure of H_2.
- I have drawn the valence structure of O_2.
- I have drawn the valence structure of N_2.

b All of these molecules are non-polar.

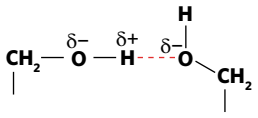
c [The molecule with the greatest bond strength will require the greatest amount of energy to decompose.[1]] [The more electrons shared in a covalent bond, the stronger the force of attraction between the atoms.[2]] [Therefore, H_2 will require the least amount of energy due to its single covalent bond, O_2 will require a greater amount of energy due to its double covalent bond, and N_2 will require the greatest amount of energy due to its triple covalent bond.[3]]

- I have identified the relationship between bond strength and the energy required to decompose a molecule.[1]
- I have identified the relationship between the number of electrons in a covalent bond and the strength of the covalent bond.[2]
- I have ordered the molecules based on increasing energy required to decompose them.[3]

d [These molecules are all non-polar.[1]] [Therefore, the only intermolecular forces present between these molecules in a mixture are dispersion forces.[2]]

- I have identified that these molecules are non-polar.[1]
- I have identified the intermolecular forces present.[2]

13 a [Serine's side chain forms an O-H covalent bond and therefore, there will be dispersion forces, permanent dipole-dipole forces and hydrogen bonds between serine residue side chains.[1]]



- I have identified the intermolecular forces between side chains of serine.[1]
- I have drawn the hydrogen bonding interaction between side chains of serine.

b [Isoleucine has a non-polar side chain.[1]] [Therefore, there will only be dispersion forces present between isoleucine side chains.[2]]

- I have identified the polarity of the isoleucine side chain.[1]
- I have identified the intermolecular forces present between isoleucine side chains.[2]

c [Covalent bonds are much stronger than any form of intermolecular bonding, including hydrogen bonds.[1]] [Therefore, the ability to form a -S-S- covalent bond between amino acid side chains is important for the structural support and strength of proteins.[2]]

- I have identified the strength of covalent bonds compared to intermolecular bonding.[1]
- I have explained the importance of this property.[2]

d [At higher temperatures, proteins have higher kinetic energies.[1]] [As a result of higher kinetic energies, intermolecular forces start to get disrupted resulting in the protein losing its structure.[2]]

- I have described the effect of temperature on intermolecular bonding.[1]
- I have linked the loss of protein structure with the loss of intermolecular forces.[2]

Key science skills questions

14 a
1. In a fumehood, set up a bunsen burner.
2. Pour a 10 mL sample to be tested in a test tube with a thermometer.
3. Hold the sample over the bunsen burner with a retort stand.
4. Once bubbles are observed, record the temperature reached by the sample.
5. Repeat for the other samples.

- I have considered the materials required.
- I have developed a sequential method that ensures the safety of the experimenter.

b [The independent variable is the identity of the sample (more specifically, the types of intermolecular forces present in the sample).[1]] [The dependent variable is the boiling point of the sample.[2]] [Possible controlled variables include the identity of the thermometer, the volume of liquid, the distance between the test tube and the bunsen burner's flame.[3]]

- I have explained that the thermometer has collected the data.[1]
- I have identified the dependent variable.[2]
- I have identified possible controlled variables.[3]

c [The data collected by the thermometer will reveal the boiling points of the different compounds.[1]][The stronger the intermolecular forces between the molecules, the greater the amount of energy that is required to break the bonds. This would result in a higher boiling point.[2]][By comparing the boiling temperatures of each substance, we can determine the relative strengths of dispersion forces (C_3H_8), permanent dipole-dipole forces (C_3H_7Cl) and hydrogen bonds (C_3H_7OH).[3]]

- ✓ ✗ I have explained the type of data collected.[1]
- ✓ ✗ I have explained the link between boiling point and strength of intermolecular forces.[2]
- ✓ ✗ I have linked my answer to the question.[3]

d [One safety hazard is the flammability of the compounds being used in the presence of the open bunsen burner flame. This could be minimised by wearing personal protective equipment and taking care when handling flammable substances.[1]][Another safety hazard is inhaling volatile organic compounds after they have been boiled. This could be minimised by undertaking the experiment in a fume hood.[2]]

- ✓ ✗ I have identified one safety concern and a solution.[1]
- ✓ ✗ I have identified another safety concern and a solution.[2]

6A Carbon materials

Theory review questions

1 B. Carbon allotropes all contain carbon atoms bonded in a different way.

2 A. Diamond's strong tetrahedral lattice structure makes it hard and rigid.

3 B. Diamond consists of a tetrahedral carbon lattice.

4 A. Graphite consists of flat layers of carbon atoms covalently bonded together. These layers are held together by dispersion forces.

5 B. The density of graphite is lower than diamond as there is space between the layers.

6 B. Graphite has layers of delocalised electrons, and therefore can conduct electricity.

7 B. The structure of graphene is equivalent to a single layer of graphite.

8 B. Buckyballs have a spherical structure, whereas nanotubes are cylindrical.

9 A. Unlike diamond, each carbon atom in a fullerene is bonded to only three carbon atoms.

10 A. The buckyball dome is held together by covalent bonds between carbon atoms.

Deconstructed questions

11 A **12** C

13 [The structure of diamond is a covalent network lattice, where each carbon atom is covalently bonded to four other carbon atoms.[1]][This is a hard and rigid structure due to the strength of these covalent bonds, causing it to maintain its structure while making a cut into the wood.[2]] [On the other hand, graphite is a covalent layer lattice, where each carbon atom is bonded to three others. The layers are held together by weak dispersion forces.[3]][Therefore, the weak intermolecular forces between layers of graphite cause a layer of graphene to be rubbed off onto the wood when graphite is rubbed against it.[4]]

- ✓ ✗ I have described the structure of diamond.[1]
- ✓ ✗ I have described the properties of diamond and its effect on the wood.[2]
- ✓ ✗ I have described the structure of graphite.[3]
- ✓ ✗ I have described the properties of graphite and its effect on the wood.[4]

Exam-style questions

Within lesson

14 **a** Carbon allotropes are molecules consisting of carbon atoms which have different atomic arrangements and properties.

b The carbons in both buckyballs and graphene take part in covalent bonding with three other carbon atoms.

c [Buckyballs have a spherical shape and therefore have a large surface area.[1]][As a result, it is able to carry medicine on its surface and in its hollow structure to areas of the body.[2]]

- ✓ ✗ I have identified a structural characteristic of buckyballs.[1]
- ✓ ✗ I have described how the structure of buckyballs assists in drug delivery.[2]

d **i** [Similarities between buckyballs and nanotubes include that they are both fullerenes, which are carbon allotropes that are arranged in the shape of a closed or partially closed cage.[1]][However, these molecules have many differences, for instance, nanotubes have a cylindrical shape, whereas buckyballs possess a spherical shape.[2]] [In addition to this, nanotubes and buckyballs have different properties. For example, nanotubes having high melting points, whereas buckyballs have low melting points.[3]]

- ✓ ✗ I have identified that nanotubes and buckyballs are both fullerenes.[1]
- ✓ ✗ I have identified the different spatial orientations of buckyballs and fullerenes.[2]
- ✓ ✗ I have described the different properties of nanotubes and buckyballs.[3]

ii [One feature of nanotubes that enables their use for reinforcement of fabrics is their strength.[1]][The strength of nanotubes comes as a result of their cylindrical shape, as well as the strong covalent bonding between the atoms.[2]][Additionally, the small size of nanotubes enables them to be used as fabric reinforcement without noticeably disrupting the texture of the fabric.[3]]

- ✓ ✗ I have identified strength as an appropriate feature.[1]
- ✓ ✗ I have described the structural features of nanotubes that provides them with strength.[2]
- ✓ ✗ I have identified the small size of nanotubes as another appropriate feature.[3]

iii [One of the features of nanotubes that enables their use as radio antennas is that they possess a high electrical conductivity.[1]] [This is a result of each carbon atom only being bonded to three other atoms, leaving one electron free for every carbon atom in the molecule and thus enabling conduction of electricity along the tube.[2]]

- ✓ ✗ I have identified high electrical conductivity as an appropriate feature.[1]
- ✓ ✗ I have described the structural features of nanotubes that enable electrical conductivity.[2]

15 [Graphite is a covalent layer lattice made from layers of carbon lattices held together by weak dispersion forces.[1]][The friction would break the weak dispersion forces between layers, leaving layers of graphite instead of individual carbon atoms as the covalent bonds between carbon atoms are too strong to be broken by friction.[2]][Therefore, the statement made by the artist is inaccurate.[3]]

- ✓ ✗ I have described the structure of graphite.[1]
- ✓ ✗ I have described the effect of friction on the structure of graphite.[2]
- ✓ ✗ I have linked my answer to the question.[3]

16 a [Due to the fact that carbon atoms in graphite bond to only three other carbons, the single unpaired electron from every carbon becomes a delocalised electron.[1]][The delocalised electrons are able to move through the layers of graphite when electricity is applied,[2]] [and therefore, graphite is able to conduct electricity through the movement of charged particles.[3]]

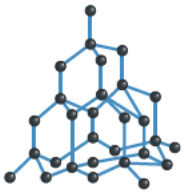
- I have described the structure of graphite.[1]
- I have identified the effect of electricity on the structure of graphite.[2]
- I have linked my answer to the question.[3]

b
- Diamonds do not conduct electricity
- Diamond is an abrasive and may damage components of the motor

c [Graphite has an extremely high melting point,[1]][and therefore the thermal energy generated in the electric motor is unlikely to interfere with the functionality of the carbon brushes.[2]][This is because the thermal energy generated would be insufficient to overcome the strength of the strong network of covalent bonds within the layers of graphite.[3]]

- I have identified the relative melting point of graphite.[1]
- I have linked my answer to the question.[2]
- I have described the effect of heat on the structure of graphite.[3]

17 a



- I have shown each carbon atom bonded to four other carbon atoms.
- I have illustrated the structure of diamond as a lattice structure.

b [An effective abrasive substance needs to be able to rub against a surface without itself being damaged.[1]][Therefore, to prevent wear, the abrasive substance must be very hard and rigid.[2]][Diamond's structure consists of a lattice of carbon atoms strongly bonded to each other through covalent bonding.[3]][As a result of this, diamond is very hard and resistant to wear, and thus a very effective abrasive.[4]]

- I have described the purpose of an abrasive.[1]
- I have identified a property of an effective abrasive.[2]
- I have described the structure of diamond.[3]
- I have identified the effectiveness of diamond as an abrasive.[4]

c [The structure of diamond consists of carbon atoms covalently bonded to each other in a lattice.[1]][As a result of this, diamond has a very hard and rigid structure, making it resistant to wear and an effective abrasive.[2]][Graphite consists of layers of carbon atoms covalently bonded to other carbon atoms in the same layer, with the layers held together by dispersion forces. While the bonds within the layers of graphite are strong, the dispersion forces between the layers are far weaker.[3]][This means the structure of graphite isn't as hard and rigid, and friction can cause layers of graphite to rub off, making it less effective than diamond when used as a drilling tool.[4]]

- I have described the structure of diamond.[1]
- I have explained how diamond's structure makes it an effective abrasive.[2]
- I have described the structure of graphite, highlighting the primary difference with diamond.[3]
- I have compared the effectiveness of both substances as a dentist tool.[4]

18 a [Graphene could be used to make the solar panels.[1]][Graphene consists of a single-layered carbon network, in which each carbon atom is bonded to three other carbon atoms, leaving a free electron. These free electrons form a delocalised sea of electrons, which permits the conduction of electricity.[2]][Furthermore, the strength and thinness of graphene makes it a very flexible substance, and therefore is able to be curved with the shape of the car.[3]]

- I have identified a suitable carbon allotrope for the solar panels.[1]
- I have explained how the structure of the carbon allotrope allows for electrical conduction.[2]
- I have described the properties of the carbon allotrope enabling its flexibility.[3]

b [The company ordered graphite instead of graphene.[1]][Since graphene is equivalent to a single layer of graphite, many of the physical properties are going to be similar,[2]][except for flexibility and thinness due to the presence of many layers in graphite.[3]]

- I have identified that the material ordered by the company.[1]
- I have explained the similarities between both carbon allotropes.[2]
- I have described the different properties of both carbon allotropes.[3]

c [The simplest way of producing graphene from graphite is through the use of sticky tape to peel off a layer off the graphite, which is equivalent to graphene.[1]][This is possible because the forces between the individual layers consist only of weak intermolecular forces, therefore would be easily separated using this method.[2]] [However, since this is still extremely difficult to carry out, and would be too time-consuming for mass production, this would not be a feasible production technique.[3]]

✓ ✗ I have described the simplest method of producing graphene.[1]

✓ ✗ I have explained how this method works.[2]

✓ ✗ I have identified the limitations of this method.[3]

Multiple lessons

19 a [Materials used in batteries need to be able to conduct electricity produced by the battery.[1]][Due to their structure, both graphene and metals contain delocalised electrons,[2]][which allows for electricity to be conducted through the material through the movement of charged particles.[3]]

✓ ✗ I have identified the key property relevant to the function of the materials in a battery.[1]

✓ ✗ I have identified the property of both substances relevant to electrical conductivity.[2]

✓ ✗ I have linked my answer to the question.[3]

b [Graphene is held together by covalent bonds between carbon atoms,[1]][whereas metals are held together by the electrostatic force of attraction between the metal cations and delocalised electrons.[2]][Considering that graphene has a higher melting point than a metal such as copper, it takes more energy to break the covalent bonds within graphene than the electrostatic forces of attraction in metals.[3]]

✓ ✗ I have described the bonding within graphene.[1]

✓ ✗ I have described the bonding within metals.[2]

✓ ✗ I have explained the difference between the strength of the structure of copper and graphene.[3]

Key science skills

20 a [As all of the values are affected in the same way,[1]][this is likely a systematic error.[2]]

✓ ✗ I have described the effect of the error.[1]

✓ ✗ I have identified the type of error.[2]

b Answers may include one of the following:

- Changing the method
- Replacing equipment
- Calibration of equipment

c The results are accurate, however they are imprecise.

6B Introduction to hydrocarbons

Theory review questions

1 B. Hydrocarbons contain carbon and have a biological origin, making them organic molecules.

2 A. Organic molecules differing by CH_2 are said to belong to the same homologous series.

3 B. Crude oil forms over millions of years.

4 A. Fractional distillation separates compounds in order of decreasing boiling point.

5 B. A reaction between a halogen and haloalkane can produce a haloalkane.

6 B. The size of a hydrocarbon chain affects its boiling point.

7 B. Unsaturated hydrocarbons have carbon to carbon double/triple bonds.

8 A. Alkenes contain carbon to carbon double bonds.

9 B. The molecule is non polar, therefore will be held together by dispersion forces.

10 B. Both molecules have the same molecular formula.

Deconstructed questions

11 C **12** D

13 [Based on the structures of both methane and octane, both structures are non polar compounds. As a result, dispersion forces will be the strongest force of attraction responsible for both methane and octane.[1]][However, octane is a much bigger molecule than methane, therefore would have stronger dispersion forces than methane.[2]][As a result, it would have a higher boiling point than methane, and therefore would be able to withstand higher temperatures.[3]][As a result, Fuel A is methane and Fuel B is octane.[4]]

✓ ✗ I have described the intermolecular bonds for both molecules.[1]

✓ ✗ I have compared the strength of the bonding between both molecules.[2]

✓ ✗ I have compared the boiling points of both molecules.[3]

✓ ✗ I have identified the name of each fuel.[4]

Exam-style questions

Within lesson

14 a [Isomers are molecules that have the same molecular formula, but can have different structures.[1]][Molecule A has a molecular formula of C_4H_6 whereas molecule B has a molecular formula of C_4H_8.[2]][Due to the fact that they both have different molecular formulas, they cannot be isomers of each other.[3]]

✓ ✗ I have defined isomers.[1]

✓ ✗ I have identified the molecular formula of each molecule.[2]

✓ ✗ I have linked my answer to the question.[3]

b [Addition reactions can involve either alkenes and alkynes.[1]][Since molecule A contains a carbon to carbon triple bond and molecule B contains a carbon to carbon double bond,[2]][both molecules would be suitable for an addition reaction.[3]]

✓ ✗ I have identified the nature of hydrocarbons taking part in addition reactions.[1]

✓ ✗ I have identified the bonding structure of the molecules given.[2]

✓ ✗ I have linked my answer to the question.[3]

c Alkane.

15 a Renewability refers to the ability for a fuel to be produced at the same rate or quicker than it can be consumed.

b [The process is referred to as fractional distillation.[1]][Fractional distillation works by heating a sample of crude oil to the point when all components are in a gaseous form, and through gradually decreasing the temperature,[2]][the crude oil is separated based on boiling points of individual hydrocarbons.[3]]

- I have identified the name process.[1]
- I have described the process.[2]
- I have identified the property used to separate the components.[3]

c [Crude oil is a fuel source derived from fossil fuels.[1]][Fossil fuels are formed over millions of years and are consumed faster than they can be produced,[2]][therefore crude oil cannot be considered as a renewable energy source.[3]]

- I have identified the source of crude oil.[1]
- I have compared the production and consumption of crude oil.[2]
- I have linked my answer to the question.[3]

Multiple lessons

16

	Saturated hydrocarbons	Unsaturated hydrocarbons
Bonding type	Carbon to carbon single bonds in the carbon chain	Carbon to carbon double or triple bond present in the carbon chain
Can participate in addition reactions	No	Yes
Contains covalent bonds	Yes	Yes

17 a **F—F**

b Halogenation/substitution

c [Considering the molecule above would take part in a substitution reaction with fluorine, this would result in one of the hydrogen atoms in the hydrocarbon chain being replaced by a fluorine atom.[1]][As such, there would be a C-F bond present. Since fluorine is more electronegative than carbon, the shared pair of electrons would be favoured towards the fluorine atom. This would result in the C-F being polar.[2]][As a result, newly formed molecule would be held together by dispersion forces and dipole-dipole interactions.[3]]

- I have described the substitution/halogenation reaction that had taken place.[1]
- I have described the nature of the carbon-fluorine bond present.[2]
- I have identified the intermolecular bonds present.[3]

18 a 40.0L = 40 000 mL

Mass of octane = $40\,000 \times 0.703 = 28120$ g

$= 2.81 \times 10^4$ g

b $n(\text{octane}) = \dfrac{2.81 \times 10^4}{(12.0 \times 8 + 1.0 \times 18)} = 246.49$ mol

$= 2.46 \times 10^2$ mol

c Energy $= (2.46 \times 10^2) \times 5460 = 1\,343\,160$ kJ

$= 1.34 \times 10^6$ kJ

19 a [Each covalent bond represents a shared pair of electrons.[1]][In a carbon to carbon double bond, there are a total of two covalent bonds.[2]][As a result, there are a total of four electrons that are being shared.[3]]

- I have identified the nature of covalent bonds.[1]
- I have identified the number of bonds in a double bond.[2]
- I have identified the total number of shared electrons.[3]

b [Due to having four valence electrons, carbon is able to take part in four covalent bonds.[1]][Since covalent bonds are strong intramolecular bonds, every additional bond between a carbon atom increases the strength of the bond.[2]][Since hydrocarbons contain hydrogen and carbon atoms, the maximal number of bonds that form between carbon atoms is a carbon to carbon triple bond, which is also the strongest carbon to carbon bond that can form in a hydrocarbon.[3]]

- I have identified the bonding of a carbon atom.[1]
- I have identified the type of bond that occurs in a hydrocarbon.[2]
- I have described the type of carbon to carbon bond with the strongest bonds.[3]

c [The single bond is the least reactive with bromine gas.[1]][This is because it is unable to participate in addition reactions, as all of its covalent electrons are occupied, and cannot be broken without a significant amount of energy and compromising the structural integrity of the compound.[2]][On the other hand, unsaturated hydrocarbons can break a single bond in a carbon to carbon double or triple bond without compromising the molecule's integrity and requires much less energy. As a result, these molecules are far more likely to take part in an addition reaction.[3]]

- I have identified the molecule least likely to react with bromine.[1]
- I have explained why saturated hydrocarbons cannot participate in addition reactions.[2]
- I have explained why unsaturated hydrocarbons can participate in addition reactions.[3]

d [Diamonds have a stronger structure than hydrocarbons.[1]][The structure of hydrocarbons consists of a linear chain of carbon atoms bound to each other, which are held together by dispersion forces, whereas diamond consists of a 3D carbon lattice held together by covalent bonds.[2]][Consequently, it more difficult to disrupt the structure of diamond compared to, wherein it is far more difficult to move, diamonds are structurally far stronger than hydrocarbons.[3]]

- I have identified that diamonds are stronger than hydrocarbons.[1]
- I have described the structures of hydrocarbons and diamonds.[2]
- I have explained why the structure of diamond is stronger than the structure of hydrocarbons.[3]

Key science skills

20 **a** The oil spill is likely to contain a large variety of different hydrocarbons with varying carbon chain lengths.

b [Hydrocarbons are not able to interact with water due the fact that hydrocarbons are predominately nonpolar whereas water is polar,[1]] [therefore is unable to mix with the water.[2]]

✓ ✗ I have described that hydrocarbons are not able to interact with water.[1]

✓ ✗ I linked my answer to the question.[2]

c A personal error (a mistake).

d The error can be accounted for by taking into account the amount of oil that was in the barrel that spilled into the water.

6C Naming simple hydrocarbons

Theory review questions

1 A. The suffix used for carbon to carbon double bonds is -ene.

2 A. The parent name prop- depicts three carbon atoms.

3 B. The suffix used for carbon to carbon single bonds is -ane.

4 B. Hydrocarbons with a carbon to carbon triple bond are alkynes.

5 B. The carbons in the parent chain of an unsaturated hydrocarbon are numbered to allow the carbon to carbon triple bond to occur at the lowest numbered carbon.

6 B. The parent naming system for a parent chain also applies for alkyl groups.

7 B. 3,3-diethylhexane contains 10 carbon atoms, 3-methyloctane contains 9 carbon atoms.

8 B. The parent chain in an organic molecule does not have to be linear.

9 B. Two of the same group is denoted by the prefix 'di'.

10 B. Numbers must be used to show the location of each branch even if they occur on the same carbon.

Deconstructed questions

11 D **12** D

13 [The longest carbon chain in this molecule is eight carbon atoms long,[1]] [and therefore the parent name should be oct- rather than hept-.[2]] [There is a methyl group on the 3rd carbon in the carbon chain rather than an ethyl group on the 2nd carbon,[3]] [and thus the prefix 3-methyl should be used.[4]] [Therefore, the correct name for this molecule is 3-methyloctane.[5]]

✓ ✗ I have identified the number of carbons in the parent chain.[1]

✓ ✗ I have identified the appropriate parent name.[2]

✓ ✗ I have identified the position of the alkyl group.[3]

✓ ✗ I have identified the appropriate notation for the alkyl group.[4]

✓ ✗ I have provided the IUPAC systematic name for this molecule.[5]

Exam-style questions

Within lesson

14 **a** **i** Pentane

ii Octane

iii Butane

b Decane, octane, heptane, hexane, propane, ethane

c [There are seven carbon atoms in 3-methylhexane.[1]] [This is because there are six carbon atoms in the parent chain of the molecule, denoted by the parent name hex-, as well as an additional carbon atom in the methyl group.[2]]

✓ ✗ I have identified the number of carbon atoms in the molecule.[1]

✓ ✗ I have described the structures of the molecule contributing to the number of carbon atoms.[2]

15 **a** **i** Ethane

ii 2-methylpropane

iii Hex-2-ene

iv 3-methylpentane

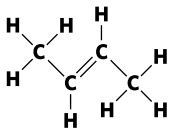
b **i** [The professor is incorrect as the longest carbon chain in the molecule is five carbon atoms long rather than four.[1]] [As a result of this, the parent name of the molecule should be pent-, rather than but-, and there should be a methyl group on carbon number 3 rather than an ethyl group located at carbon number 2.[2]]

✓ ✗ I have identified the error in the parent name of the molecule.[1]

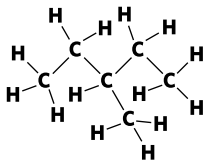
✓ ✗ I have described the changes that should be made to make the name correct.[2]

ii 3-methylpentane

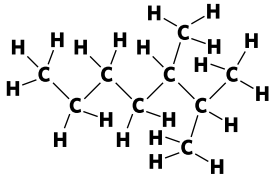
16 **a** **i**



ii



iii



b [The name of this molecule does not follow correct IUPAC nomenclature.[1]] [When naming a molecule, side groups must be counted such that they are the shortest distance away from the end of the main carbon chain, regardless of the spatial orientation of the molecule. This molecule should therefore be 3-methyl not 6-methyl.[2]] [Since the naming of the molecule is incorrect, this molecule cannot exist.[3]]

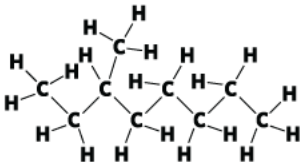
✓ ✗ I have identified that this molecule does not follow IUPAC nomenclature.[1]

✓ ✗ I have explained why this molecule's naming is incorrect.[2]

✓ ✗ I have linked my answer to the question.[3]

c 3-methyloctane.

d



Multiple lessons

17 a Isomers are molecules with the same molecular formula, but with different atomic organisation.

b [1,1-dimethylbutane could not exist as one of the methyl groups on the first carbon atom would contribute to the parent carbon chain.[1]] [Instead, this molecule would be named 2-methylpentane.[2]]

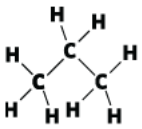
✓ ✗ I have explained why 1,1-dimethylbutane could not exist.[1]

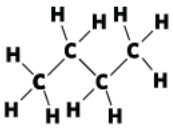
✓ ✗ I have identified the correct IUPAC systematic name of the molecule.[2]

c Hexane

18 a An organic molecule is a carbon-based molecule (with some exceptions such as carbonates).

b i [The term homologous series refers to a number of molecules that all have similar structures and properties.[1]] [Hydrocarbons are a homologous series of carbon-based molecules, each centred around a carbon chain and differing from adjacent molecules in the series by a CH_2 group.[2]] [As shown in the diagrams, the hydrocarbons shown both consist of carbon and hydrogen, and have a central carbon chain. They also differ from each other by a CH_2 group, as the molecule on the right possesses one more.[3]]





✓ ✗ I have defined the term homologous series.[1]

✓ ✗ I have described hydrocarbons in a homologous series.[2]

✓ ✗ I have referred to at least two diagrams demonstrating how the structures of hydrocarbons are a homologous series.[3]

✓ ✗ I have included at least two diagrams.[4]

ii [IUPAC nomenclature of hydrocarbons reflects a homologous series as they are all based around a common structural feature.[1]] [This is reflected in the suffix parent name, which describes the type of bonding in the parent carbon chain.[2]]

✓ ✗ I have described how IUPAC nomenclature reflects that hydrocarbons are a homologous series.[1]

✓ ✗ I have identified the assumption of a homologous series in the parent names of hydrocarbons.[2]

c [The IUPAC systematic name of this molecule is 2-methyl-4-propyloctane.[1]] [The longest carbon chain in this molecule contains eight carbon atoms and consists of only single carbon to carbon bonds. Therefore the parent name oct- is used.[2]] [Additionally, there is a methyl group on the 2nd carbon atom, counting from the left, and a propyl group on the 4th carbon atom. Methyl is named before propyl as alkyl groups must be listed in alphabetical order.[3]]

✓ ✗ I have identified the IUPAC systematic name of the molecule.[1]

✓ ✗ I have explained the parent name of this molecule.[2]

✓ ✗ I have explained the naming of alkyl groups in this molecule.[3]

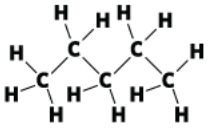
19 a i Alkane

ii Alkane

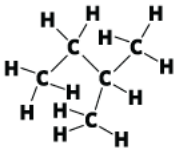
iii Alkyne

iv Alkene

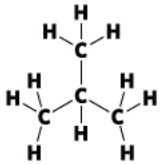
b



Pentane



2-methylbutane



2,2-dimethylpropane

Key science skills

20 a i [The results are inaccurate as they are not close to the true value,[1]] [however they are precise as the measured values are within close proximity of each other.[2]]

✓ ✗ I have commented on the accuracy of the results.[1]

✓ ✗ I have commented on the precision of the results.[2]

ii [The results indicate a systematic error.[1]] [This is because all of the results have been affected to the same degree, since they are precise, which indicates that there may be an error with the procedure or the instruments being used, rather than a personal error.[2]]

✓ ✗ I have identified that the most likely type of error is a systematic error.[1]

✓ ✗ I have explained why the results are likely the result of a systematic error.[2]

b Answers may include:

- Tie hair.
- Use of instruments to maintain distance from flame e.g. crucible tongs.
- Use a fume hood to protect from hazardous fumes.
- Use safety flame when flame is not in use.

6D Naming complex hydrocarbons

Theory review questions

1 B. The parent chain is numbered based on the position of carbon to carbon double double/triple bonds first.

2 A. Both molecules have the same parent chain, and a methyl group located at the same numbered carbon.

3 A. Carbon to carbon double bonds are prioritised when numbering carbons in the parent chain.

4 B. The carbon to carbon double bond is found on carbon number 1.

5 A. The IUPAC naming system will always present carbon to carbon double/triple bonds at the lowest numbered carbon.

6 B. Electron dot formulas show the valence electrons of each atom.

7 B. Skeletal structures do not show chemical symbols of atoms associated in the parent chain.

8 A. Each point in the structure represents a carbon atom.

9 A. The structure contains four carbons in the parent chain.

10 A. There is a total of five carbons in the molecule.

Deconstructed questions

11 B **12** B

13 3-ethylpent-2-ene

Exam-style questions

Within lesson

14 D **15** B

16 **a** 2-methylpent-2-ene

b 3,3-dimethylbut-1-ene

c 3,3-dimethylpent-1-ene

d 2-methyloct-3-ene

Multiple lessons

17 **a** Isomers

b **i** C_7H_{14}

ii Molecule A: 2-methylhex-3-ene

Molecule B: 3-methylhex-3-ene

18 **a** Molecular formula of the compound: C_8H_{16}

$M_r(C_8H_{16}) = 12.0 \times 8 + 1.0 \times 16 = 112\ g\ mol^{-1}$

b 3-ethylhex-3-ene

c [Unsaturated hydrocarbon contain at least one carbon to carbon double or triple bond in the parent chain, whereas saturated hydrocarbons contain only single carbon to carbon single bonds in the parent chain.[1]][Due to the carbon to carbon double bond located at carbon 3 in the parent chain,[2]][this molecule is considered an unsaturated hydrocarbon.[3]]

✓ ✗ I have compared the definitions of saturated and unsaturated hydrocarbons.[1]

✓ ✗ I identified a key characteristic of the structural formula.[2]

✓ ✗ I have linked my answer to the question.[3]

19 **a** [Due to the small difference in electronegativity between carbon and hydrogen atoms, the bonds occurring between carbon and hydrogen in the hydrocarbon are non-polar.[1]][As a result, the molecule shown is a non-polar compound, and therefore the bonds responsible for holding these molecules together are weak dispersion forces.[2]][As a result, to change phase of the substance made up of these compounds, the only intermolecular bonds that need to be broken are dispersion forces.[3]]

✓ ✗ I have described the bonding between carbon and hydrogen atoms.[1]

✓ ✗ I have described the non-polar nature of the compound.[2]

✓ ✗ I have linked my answer to the question.[3]

b $CH_2 = CHCH_2CH(CH_3)CH_2CH_3$

c $m(C_7H_{14}) = 0.122 \times (12.0 \times 7 + 1.0 \times 14) = 12.0\ g$

Key science skills

20 **a** Boiling point of each compound

b [Saturated hydrocarbons contain only carbon to carbon single bonds in their parent chain.[1]][Of the compounds being tested, only propane contains only carbon to carbon single bonds in the parent chain,[2]][and as a result this is the only one that can be considered a saturated hydrocarbon.[3]]

✓ ✗ I have defined saturated hydrocarbons.[1]

✓ ✗ I have identified the compound that fits the criteria.[2]

✓ ✗ I have linked my answer to the question.[3]

c [The thermometer in the experiment is responsible for measuring the temperature of the water in the beaker.[1]][It is assumed that the temperature measured is the temperature of the environment in which the organic compound in the test tube is subjected to.[2]][However, as the test tube is touching the bottom of the beaker, it is actually subjected to higher temperatures as it is closest to the heating pad of the hot plate.[3]][As the sample in the test tube is subjected to higher temperatures than the temperature of the water, it is possible that the temperature recorded for the boiling point of the organic compound is lower than the true value,[4]][and as a result, the results may be inaccurate.[5]]

✓ ✗ I have identified the function of the thermometer in the experiment.[1]

✓ ✗ I have identified the assumption in the experiment.[2]

✓ ✗ I have described the effect of the positioning of the test tube.[3]

✓ ✗ I have described the effect of the error on the data measured.[4]

✓ ✗ I have linked my answer to the question.[5]

6E Functional groups - alcohols and carboxylic acids

Theory review questions

1 A. A hydroxyl group is a functional group consisting of an oxygen and a hydrogen atom (-OH).

2 A. The prefix 'ethan' denotes a two-carbon saturated hydrocarbon, and the suffix '-ol' denotes an alcohol.

3 A. The parent name propan- is used to indicate three-carbon atoms are in the saturated carbon chain, and the suffix -oic acid indicates this is a carboxylic acid.

4 B. The -OH group in the carboxyl group is able to take part in hydrogen bonding.

5 B. Structural isomers have the same molecular formula but different spatial arrangements.

6 B. The influence of functional groups decreases as the carbon chain length increases.

7 A. When numbering functional groups, the smallest number must always be used.

8 B. Alcohols are commonly used as a fuel.

9 B. Carboxylic acids have a higher boiling point than alcohols as the carbonyl group contributes to the polarity of the molecule.

10 A. Dimers are carboxylic acids with hydrogen bonding between their carboxyl groups.

Deconstructed questions

11 A **12** D

13 [Butan-2-ol is named correctly, however the molecule labelled butan-4-ol should be named butan-1-ol instead.[1]][This is because functional groups should always be named on the smallest numbered carbon atom.[2]] [When counting from left to right, the hydroxyl group appears to be on the 4th carbon atom, however counting from right to left the hydroxyl group is on the 1st carbon atom. Therefore, in order to have the functional group on the lowest numbered carbon atom, this molecule will be named butan-1-ol rather than butan-4-ol.[3]]

- ✓ ✗ I have identified the correct names of the molecules presented.[1]
- ✓ ✗ I have explained that functional groups are named on the lowest numbered carbon atom.[2]
- ✓ ✗ I have explained why butan-4-ol should be butan-1-ol.[3]

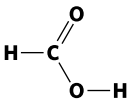
Exam-style questions

Within lesson

14 a [Methanoic acid is a carboxylic acid,[1]][and therefore possesses a carboxyl functional group.[2]]

- ✓ ✗ I have identified that methanoic acid is a carboxylic acid.[1]
- ✓ ✗ I have identified that the functional group found in a carboxylic acid.[2]

b



c Methanoic acid is only capable of causing bodily harm at high concentrations.

d i The molecule is an alcohol.

ii Methanol

iii



15 a [The addition of a hydroxyl group makes the hydrocarbon polar.[1]] [This is due to a difference in electronegativity between the oxygen and hydrogen atoms in the hydroxyl functional group, making it a polar bond and therefore a polar functional group.[2]][Consequently, hydrocarbons with hydroxyl groups are able to form stronger dipole-dipole interactions or potentially hydrogen bonds with other molecules.[3]]

- ✓ ✗ I have identified the effect of a hydroxyl group on the polarity of a compound.[1]
- ✓ ✗ I have explained why hydroxyl groups are polar functional groups.[2]
- ✓ ✗ I have described the intermolecular forces that will be seen in hydrocarbons with hydroxyl groups.[3]

b [The stronger intermolecular forces between alcohols will result in a higher boiling point than in hydrocarbons without functional groups.[1]] [This is because more energy is needed to break the hydrogen bonds between alcohol molecules in comparison to the dispersion forces between alkane molecules.[2]]

- ✓ ✗ I have identified the compound with a higher boiling point.[1]
- ✓ ✗ I have explained that the increased boiling point relates to the increased amount of energy needed to break intermolecular bonds.[2]

c i Carboxyl group

ii [A dimer is a molecular complex consisting of two carboxylic acids bonded together.[1]][The carboxylic acids take part in hydrogen bonds between the -C=O and -OH groups of each of the molecules.[2]]

- ✓ ✗ I have identified what a dimer is.[1]
- ✓ ✗ I have described the intermolecular bonding arrangement in a dimer.[2]

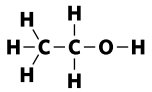
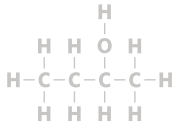
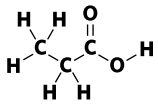
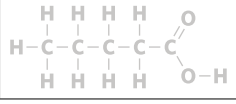
Multiple lessons

16 a i Methanol **ii** Butan-1-ol

iii Ethanoic acid **iv** Pentan-3-ol

v Heptanoic acid

b

IUPAC name	Semi-structural formula	Structural formula
Ethanol	CH_3CH_2OH	
Butan-2-ol	$CH_3CH_2CH(OH)CH_3$	
Propanoic acid	CH_3CH_2COOH	
Pentanoic acid	$CH_3CH_2CH_2CH_2COOH$	

17 a Ethanol, butanoic acid, hexan-3-ol

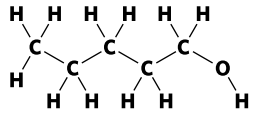
b

Name	Molecular formula	Semi-structural formula
Ethanol	C_2H_6O	CH_3CH_2OH
Butanoic acid	$C_4H_8O_2$	$CH_3CH_2CH_2COOH$
Hexan-3-ol	$C_6H_{14}O$	$CH_3CH_2CH_2CH(OH)CH_2CH_3$

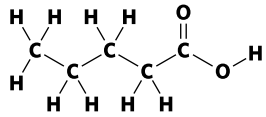
c [Carboxylic acids are able to be mixed in water due to the presence of a carboxyl group,[1]][which enables it to form strong hydrogen bonds with water molecules.[2]][The intermolecular forces formed between carboxylic acid and water are able to overcome the forces between water molecules, allowing it to be mixed in the water.[3]]

- I have identified the role of the carboxyl group to the molecule's ability to mix in water.[1]
- I have explained how the carboxyl group contributes to mixing in water.[2]
- I have explained the relationship between hydrogen bonding and water solubility.[3]

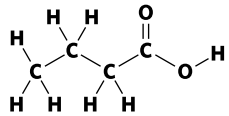
18 a i



ii



iii



b [Pentanoic acid would be expected to have the highest boiling point of these molecules.[1]][This is because it is able to form strong hydrogen bonds with other pentanoic acid molecules due to the hydroxyl group, and, additionally, the presence of a carbonyl group also makes it more polar than the other molecules. Furthermore, carboxylic acids are able to form dimers, which consist of two molecules bonded together by strong intermolecular forces, which further increase the boiling point of the substance.[2]][By comparison, the molecule butanoic acid has weaker intermolecular forces due to the shorter carbon chain,[3]][and the hydroxyl functional group in pentan-1-ol would not be able to form intermolecular bonds as strong as those able to be formed by the carboxyl group.[4]][Consequently, the intermolecular forces between pentanoic acid molecules are going to be the strongest, and therefore it will have the highest boiling point.[5]]

- I have identified the molecule with the highest boiling point.[1]
- I have identified the features of pentanoic acid that contribute to intermolecular bonding.[2]
- I have explained the impact of carbon chain length on boiling point.[3]
- I have compared the bonding between a hydroxyl group and a carbonyl group.[4]
- I have explained the link between strong intermolecular forces and boiling point.[5]

c Pentanoic acid

d i [These molecules all have a higher boiling point than their alkane counterpart due to the presence of a functional group.[1]][These functional groups have greater polarity than the non-polar bonds in the alkanes, making the overall molecule more polar and thus able to form stronger dipole-dipole interactions than the dispersion forces in alkanes,[2]][which produces a higher boiling point as more energy is required to break the intermolecular forces.[3]]

- I have identified the feature which provides these molecules with a higher boiling point than their alkane counterparts.[1]
- I have explained how functional groups contribute to the intermolecular interactions of molecules.[2]
- I have identified that these intermolecular forces increase the boiling point.[3]

ii [The difference between the boiling points of an alcohol and its alkane counterpart decreases as the carbon chain length increases because of the reduced influence of the hydroxyl group.[1]][As the length of the carbon chain increases, the dispersion forces in both molecules become stronger[2]][and, since the non-polar portion of the molecule increases, the polarity of the molecule decreases, thus decreasing the ability of the alcohols to form dipole-dipole bonds.[3]][These factors together result in the boiling points becoming increasingly similar as the carbon chain length increases.[4]]

- ✓ ✗ I have identified that the hydroxyl group gradually exerts less influence on intermolecular forces as the carbon chain increases.[1]
- ✓ ✗ I have explained that dispersion forces increase in both molecules.[2]
- ✓ ✗ I have explained that the capacity for dipole-dipole interactions is reduced.[3]
- ✓ ✗ I have described how these factors result in the boiling points coming closer together.[4]

19 a [Octane is a non-polar alkane without any functional groups, and therefore it can only form weak dispersion forces.[1]][In contrast, the presence of a hydroxyl group in octan-1-ol enables it to form strong hydrogen bonds with other octan-1-ol molecules in conjunction with dispersion forces and permanent dipole-dipole forces.[2]] [Consequently, the intermolecular forces between octan-1-ol molecules are going to require much more energy to overcome than those between octane molecules, and therefore octan-1-ol will have a higher boiling point.[3]]

- ✓ ✗ I have described the intermolecular forces between octane molecules.[1]
- ✓ ✗ I have described the intermolecular forces between octan-1-ol molecules.[2]
- ✓ ✗ I have explained that stronger intermolecular forces require more energy to overcome.[3]

b i Renewable energy resources will never run out and can produce energy indefinitely as they can be replenished at a rate faster than or equal to the rate at which they are consumed.

ii [The presence of a hydroxyl group enables ethanol to form strong hydrogen bonds with water molecules.[1]][As a consequence, ethanol in fuel may have some water molecules bound to it via these intermolecular forces, introducing additional water molecules into the engine.[2]][As a result, an increase in ethanol concentration would result in an increase in the amount of water present in the engine.[3]]

- ✓ ✗ I have identified ethanol's ability to form hydrogen bonds with water.[1]
- ✓ ✗ I have explained the impact of the interaction between ethanol and water on fuels.[2]
- ✓ ✗ I have linked my answer to the question.[3]

Key science skills

20 a Answers may include:

- Keep test tube in a fume hood to protect from fumes
- Tie hair back
- Use gloves when handling chemical substances
- Use goggles to protect eyes

b [The independent variable in this experiment is the type of organic substance used.[1]][This is because the independent variable is the variable changed by the experimenter, and the experiment involves exposing alcohols and carboxylic acids to a flame for comparison.[2]]

- ✓ ✗ I have identified the independent variable in this experiment.[1]
- ✓ ✗ I have explained why the type of substance is the independent variable.[2]

c The dependent variable is the boiling point.

d Controlled variables may include:

- Amount of the substances used
- Amount of heat from the bunsen burner
- Position relative to flame
- Type of test tube
- Equipment used to measure boiling point

6F Functional groups - esters

Theory review questions

1 B. The ester functional group consists only of carbon and oxygen.

2 B. The ester functional group (COO) contains one carbon-oxygen double covalent bond and one carbon-oxygen single covalent bond.

3 A. The formation of an ester produces water (condensation) whereas the breakdown of an ester requires water (hydrolysis).

4 B. An alcohol and carboxylic acid react to produce an ester in an esterification reaction.

5 B. Esterification reactions require a concentrated sulfuric acid catalyst to increase the rate of reaction so that it can be observed.

6 B. Esters are responsible for the smell of many fruits.

7 B. The alcohol from which an ester is derived is associated with the single C-O covalent bond.

8 A. The carboxylic acid from which an ester is derived is associated with the carbonyl C=O bond.

9 B. Esters can be hydrogen bond acceptors (i.e. hydrogen bond with a substance like water) because they have highly electronegative oxygen atoms.

10 A. Carboxylic acids can form hydrogen bonds with other carboxylic acid molecules due to the presence of the hydroxyl group which can function as a hydrogen bond donor and acceptor.

Deconstructed questions

11 C **12** C

13 [Esterification reactions are condensation reactions which means that one of the products is water. Therefore, as the reaction proceeds, there is more water being produced, resulting in the increase in volume of water.[1]] [Esters also have a sweet, fruity smell and so therefore, when they are produced, a fruity smell will emerge from the test tube.[2]]

- ✓ ✗ I have explained the increasing volume of water in the test tube.[1]
- ✓ ✗ I have explained the fruity smell coming from the test tube.[2]

Exam-style questions

Within lesson

14 D

15 **a** **i** Butyl methanoate

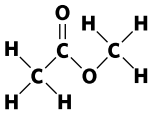
ii Propyl pentanoate

iii Hexyl butanoate

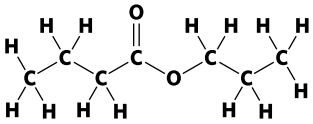
b **i** Ethyl pentanoate **ii** Propyl propanoate

iii Methyl hexanoate **iv** Methyl propanoate

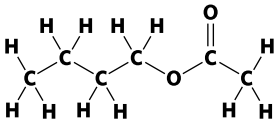
c **i**



ii



iii



Multiple lessons

16 **a** [Diesel consists of alkanes which are hydrocarbons with no functional groups present.[1]][Biodiesel, since it consists of esters, has the ester -COO- functional group present.[2]]

I have identified the functional groups present in diesel.[1]

I have identified the functional groups present in biodiesel.[2]

b [Since diesel consists of alkanes with no functional groups present, the only intermolecular force present will be dispersion forces.[1]][Biodiesel, with the polar ester functional group, will have dispersion forces as well as permanent dipole-dipole forces.[2]][Since melting point depends on the strength of intermolecular forces, biodiesel will have a higher melting point due to the strength of the permanent dipole-dipole forces.[3]][In cold conditions, it is better to use diesel as it is more likely to be in a liquid state which is necessary for flow along fuel lines.[4]]

I have identified the intermolecular forces present in diesel.[1]

I have identified the intermolecular forces present in biodiesel.[2]

I have linked the strength of intermolecular forces to boiling point.[3]

I have justified which fuel choice is better in cold conditions.[4]

c [Since hygroscopicity is the property of fuels to absorb water, the fuel which is able to bond with water will be the most hygroscopic.[1]][Since diesel has no polar functional groups present, therefore it will not interact with water.[2]][However, biodiesel can form hydrogen bonds with water due to the polar ester functional group, resulting in a greater hygroscopicity.[3]][In humid conditions, since there is a lot of water present in the air, it is better to use diesel to avoid the quality of the fuel decreasing as it absorbs water.[4]]

I have linked hygroscopicity to the property of a substance to bond with water.[1]

I have identified the intermolecular forces possible between diesel and water.[2]

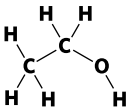
I have identified the intermolecular forces possible between biodiesel and water.[3]

I have justified which fuel choice is better in humid conditions.[4]

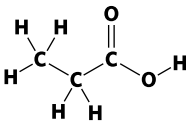
17 **a** Alcohol + Carboxylic acid $\xrightarrow{H_2SO_4}$ Ester + Water

b Esterification reactions are condensation reactions as they produce a water molecule.

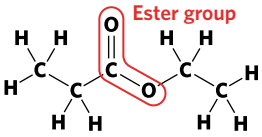
c



d



e



f Ethyl propanoate

18 **a** **i** Ester: Methyl propanoate, Alcohol: Methanol, Carboxylic acid: Propanoic acid

ii Ester: Propyl pentanoate, Alcohol: Propan-1-ol, Carboxylic acid: Pentanoic acid

iii Ester: Propyl ethanoate, Alcohol: Propan-1-ol, Carboxylic acid: Ethanoic acid

b [Water is a product of an esterification reaction.[1]][Therefore, in order to reverse this reaction, water is needed in a hydrolysis reaction to reform the carboxyl and hydroxyl groups of the carboxylic acid and alcohol respectively.[2]]

I have identified water as a product of esterification reactions.[1]

I have explained why water is needed for hydrolysis reactions.[2]

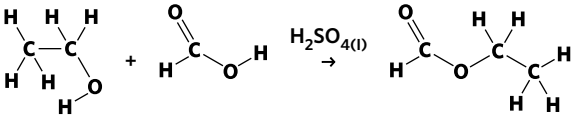
c [Concentrated sulfuric acid acts as a catalyst and is not consumed in esterification reactions.[1]][Its purpose is to increase the rate of the chemical reaction so that it occurs in a timely manner.[2]]

I have identified sulfuric acid as a catalyst.[1]

I have explained the purpose of a catalyst.[2]

19 **a** Esterification reaction

b



c Ethyl methanoate and water

d [A byproduct is a product of a chemical reaction that is not desired.[1]] [In this case, since the main product of an esterification reaction is the ester, water is the byproduct.[2]]

- ✓ ✗ I have defined a byproduct.[1]
- ✓ ✗ I have explained why water is the byproduct.[2]

e A fruity, sweet smell coming from the test tube.

f [Since all three molecules have similar molar masses, the strength of dispersion forces will be similar.[1]][However, ethanoic acid is a carboxylic acid and therefore, the carboxyl group can form strong hydrogen bond interactions with another carboxyl group in a dimer, leading to the highest strength of intermolecular bonding out of the three molecules.[2]][Ethanol is an alcohol and therefore has a hydroxyl group which can also hydrogen bond with another hydroxyl group, albeit less strongly than the carboxyl functional group.[3]] [Finally, methyl methanoate contains an ester functional group which cannot hydrogen bond with another ester functional group, but can form permanent dipole-dipole forces, leading to weaker intermolecular forces.[4]][Since boiling point depends on the strength of the intermolecular forces, the boiling point of ethanoic acid>ethanol>methyl methanoate because it takes more energy to disrupt the intermolecular forces in ethanoic acid, compared to ethanol and then methyl methanoate.[5]][This property could be used to heat the mixture to the boiling point of the ester but not the alcohol or carboxylic acid. Therefore, the ester will evaporate but not the carboxylic acid or alcohol.[6]]

- ✓ ✗ I have identified the similarities between all molecules.[1]
- ✓ ✗ I have described the intermolecular forces of ethanoic acid.[2]
- ✓ ✗ I have described the intermolecular forces of ethanol.[3]
- ✓ ✗ I have described the intermolecular forces of methyl methanoate.[4]
- ✓ ✗ I have linked the strength of intermolecular forces to boiling point.[5]
- ✓ ✗ I have explained how this property could be used to extract the ester.[6]

Key science skills

20 [The student could waft the vapour towards her nose rather than directly smelling the tube.[1]][Directly smelling the test tube is dangerous because organic compounds can be toxic when inhaled in large quantities.[2]]

- ✓ ✗ I have suggested an alternate method of smelling the ester.[1]
- ✓ ✗ I have explained why her original experimental method is dangerous.[2]

b i Systematic

ii An error that consistently affects the results by the same degree

iii The error could be corrected by determining the amount that the scale is off by, and then deducting or adding that amount to the final results to counter the incorrect calibration.

6G Polymers

Theory review questions

1 A. Polymers consist of monomers in a chain with some branches.

2 B. Unsaturated hydrocarbons can be monomers.

3 A. Addition polymerisation is the formation of a polymer through adding together monomers.

4 A. One representation of a polymer is square brackets with *n* in subscript after it.

5 B. Low density polyethene is produced at higher pressure than high density polyethene.

6 A. Thermoplastic polymers have weaker intermolecular bonds which enable remoulding.

7 B. Covalent intermolecular bonding causes hardness and degradation when heated.

8 B. Thermoplastic polymers can have intermolecular bonds such as hydrogen bonds, dipole-dipole bonds or weak dispersion forces.

9 A. Cross-links are covalent bonds between polymer chains.

10 A. Elastomers have cross-links less regularly than thermosetting polymers and therefore are elastic rather than hard.

Deconstructed questions

11 A **12** D

13 [Polymers are composed of monomers bound together in a chain.[1]] [The monomers are unsaturated hydrocarbons which are bound together through addition polymerisation, breaking one bond in their structure in order to form a bond with another monomer.[2]]

- ✓ ✗ I have identified the general structure of polymers.[1]
- ✓ ✗ I have explained how polymers are formed.[2]

Exam-style questions

Within lesson

14



Ethene Polyethene

15 B **16** D **17** A

18 a Addition polymerisation

b i [At a higher pressure setting, the polymer produced would be low density polyethene.[1]][Due to the presence of branching in LDPE, this type of polymer has weaker dispersion forces between polymer chains, which causes the molecule to possess properties such as being relatively soft and having a low melting point.[2]]

- ✓ ✗ I have identified the polymer produced at a higher pressure.[1]
- ✓ ✗ I have explained the properties of LDPE.[2]

ii [At a lower pressure setting, the polymer produced would be high density polyethene.[1]][Since HDPE does not form branches in its structure, HDPE chains are able to pack more closely together than LDPE, increasing the strength of dispersion forces between chains and leading to properties such as a higher melting point and hardness.[2]]

- I have identified the polymer produced at a lower pressure.[1]
- I have explained the properties of HDPE.[2]

c i [Grant has produced a thermosetting polymer.[1]][This polymer decomposes when heated because the intermolecular bonds are of similar strength to the intramolecular bonds, and therefore when the polymer is heated both bonds are equally likely to break.[2]] [This means that the substance cannot be remoulded, which relies on overcoming intermolecular bonds to move polymers, and instead decomposes the structures of individual polymers.[3]]

- I have identified that Grant produced a thermosetting polymer.[1]
- I have explained the similarity in strength between intermolecular and intramolecular bonds.[2]
- I have explained how the bonding arrangement leads to decomposition on heating.[3]

ii Answers may include:

- Hard
- Rigid
- Heat resistant

Multiple lessons

19 a i Can be used

ii [Methane cannot be used to form a polymer.[1]][This is because monomers must be unsaturated to take part in an addition polymerisation reaction, as this reaction involves removing one bond to form a bond with another monomer.[2]]

- I have identified that methane cannot be used to form a polymer.[1]
- I have explained that addition polymerisation involves removing a bond to form another bond.[2]

iii Can be used

iv Can be used

b i Thermoplastic polymers

ii [Thermoplastic polymers possess intermolecular bonds such as hydrogen, dipole-dipole and weak dispersion forces, whereas elastomers possess occasional strong covalent bonds holding the polymer chains together.[1]][Since these strong intermolecular bonds are only found occasionally in elastomers, these polymers are able to be reshaped, as desired by Edward.[2]][However, since elastomers still possess strong cross-links, these polymers are elastic and will not hold a new shape.[3]]

- I have identified the structural differences between thermoplastic polymers and elastomers.[1]
- I have explained that elastomers can be shaped.[2]
- I have described the elastic property of elastomers.[3]

Key science skills

20 a Bendability of the plastic

b Qualitative data

c [The bendability of the plastic is a subjective judgement made by the experimenter.[1]][As a result, the classification of 'bendability' will be different between different experimenters.[2]][Therefore, the results from the experiment would not be very reproducible as the results would vary depending on the experimenter.[3]]

- I have identified that 'bendability' is subjective.[1]
- I have described the impact of this factor on the results of the experiment.[2]
- I have linked my answer to the question.[3]

6H Modifying polymers

Theory review questions

1 B. A copolymer is a single polymer with at least two different types of monomers.

2 B. Non-polar polymers possess properties such as being non-stick, heat resistant and low friction.

3 A. Crystalline regions enable tight packing of chains preventing light from transmitting through.

4 B. The 'iso' prefix in isotactic indicates that all functional groups are on the same side.

5 A. Plasticisers are small particles that are added to polymers in order to increase the distance between polymer chains.

6 B. Some polymers can be recycled through reprocessing and depolymerisation.

7 A. Reprocessing and depolymerisation are techniques used for recycling polymers.

8 A. Increasing the chain length of polymers increases the strength of dispersion forces between chains.

9 A. Branching prevents tight-packing of polymers, leading to more amorphous regions.

10 B. Polar monomers result in strong intermolecular bonding between polymer chains.

Deconstructed questions

11 C **12** A

13 Advantages of polymers include the following:

- Able to be produced from renewable sources
- Readily modifiable
- High strength material
- Often recyclable

Disadvantages of polymers include the following:

- Primarily produced from fossil fuels
- Often unable to be broken down by microorganisms
- Toxic gases produced when burned
- Thermosetting polymers are often not recyclable

Exam-style questions

Within lesson

14 D

15 **a** [Polymer 1 has a shorter chain length than polymer 2, which means that the former will have weaker intermolecular bonding than the latter.[1]][Therefore, polymer 1 will be softer and more flexible than polymer 2.[2]]

✓ ✗ I have described how shorter chain length affects intermolecular bonding.[1]

✓ ✗ I have described how shorter chain length affects polymer softness and flexibility.[2]

b [Polymer 1 has more branching than polymer 2, which means that polymer 1 will be prevented from packing together tightly, reducing the strength of intermolecular bonding and thus making polymer 1 a softer and more flexible material.[1]][Conversely, the reduced branching in polymer 2 enables the polymer chains to be packed together tightly, producing strong intermolecular bonds resulting in a harder and more rigid material.[2]]

✓ ✗ I have explained how polymer 1's increased branching affects its properties.[1]

✓ ✗ I have explained how polymer 2's decreased branching affects its properties.[2]

c [A possible use of the polymer would be pipes.[1]][Increased chain length and minimal branching results in strong intermolecular bonding.[2]][The increased chain length strengthens bonding since many electrons are able to participate in bonds. Furthermore, the small amount of branching allows for increased packing between chains which also contributes to strong intermolecular bonding.[3]] [As a result of the strong bonding, a polymer which is strong, hard and rigid will be useful in pipes.[4]]

✓ ✗ I have identified a potential use of the polymer.[1]

✓ ✗ I have identified that large chain length and minimal branching leads to strong bonding.[2]

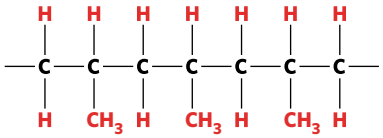
✓ ✗ I have explained how chain length and the extent of branching affects the strength of intermolecular forces.[3]

✓ ✗ I have identified how bonding strength affects properties of the polymer and linked this back to the example provided.[4]

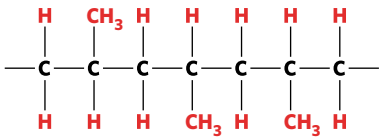
Note: Students may have different responses. Please check with your teacher.

16 **a** The unique feature of polypropene is the presence of methyl groups.

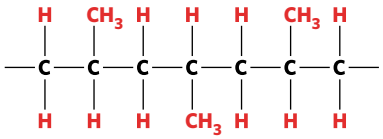
b



Isotactic polypropene



Atactic polypropene



Syndiotactic polypropene

c **i** [As the polymer is soft and flexible, it is most likely that it consists of atactic polypropene.[1]][This is because atactic polypropene has methyl groups randomly distributed along its chain, which prevents strong intermolecular bonding and tight packing of chains. This reduces the amount of crystalline regions and the overall strength of the polymer.[2]]

✓ ✗ I have identified that the material most likely consists of atactic polypropene.[1]

✓ ✗ I have explained why atactic polypropene is not as strong as other forms of polypropene.[2]

ii [Ludwig would not be able to ascertain the type of polymer by merely assessing the properties of the material.[1]][This is because determining the properties of a material such as softness or flexibility is highly subjective, and therefore may not be accurate.[2]]

✓ ✗ I have identified that Ludwig cannot be certain about the type of polymer present.[1]

✓ ✗ I have explained that determining softness and flexibility is highly subjective.[2]

Multiple lessons

17 **a** [The primary source of polymers is petroleum, which is a fossil fuel.[1]] [Another source is sugarcane.[2]][Since petroleum is a fossil fuel, it is not renewable. However polymers produced from sugarcane are renewable.[3]]

✓ ✗ I have identified the primary source of polymers.[1]

✓ ✗ I have identified another source of polymers.[2]

✓ ✗ I have compared the renewability of polymers from petroleum and sugarcane.[3]

b **i** Biodegradable polymers are able to be broken down into smaller components by microorganisms meaning they do not accumulate in the environment.

ii Answers may include:

- Produce polymers from renewable sources such as corn and starch.
- Adding additives to polymers to enable biodegradation over time.

c i [One method of polymer recycling is reprocessing, which involves breaking down polymers through shredding, melting and reshaping into a new product.[1]][The other process of recycling is depolymerisation where polymers are broken down into monomers, which are used to produce new polymers.[2]]

- ✓ ✗ I have described the process of polymer reprocessing.[1]
- ✓ ✗ I have described the process of depolymerisation.[2]

ii [The primary drawback of reprocessing is that it can only be used for thermoplastic polymers.[1]][This is because the strong intermolecular covalent bonds between chains of thermosetting polymers prevent them from being heated and remoulded.[2]] [The main limitation of depolymerisation is that it requires a significant amount of energy and has a low yield.[3]][For these reasons, depolymerisation is often not considered an economically viable option.[4]]

- ✓ ✗ I have identified the primary drawback of reprocessing.[1]
- ✓ ✗ I have explained the effect of strong intermolecular covalent bonds in thermosetting polymers on the ability of the polymer to be remoulded.[2]
- ✓ ✗ I have identified the primary drawback of depolymerisation.[3]
- ✓ ✗ I have identified the economic viability of depolymerisation.[4]

18 a [Copolymers are polymers that consist of at least two different types of monomers.[1]][The polymer that is produced will exhibit properties depending on the monomers present.[2]][Consequently, by changing the type and quantity of monomers present, the characteristics of the polymer can be altered.[3]]

- ✓ ✗ I have identified the structure of copolymers.[1]
- ✓ ✗ I have identified the impact of monomers on the properties of copolymers.[2]
- ✓ ✗ I have explained how the properties of a polymer can be altered.[3]

b i A conductive polymer is one that is able to conduct electricity due to the movement of delocalised electrons.

ii H—C≡C—H

iii [Polyethene is a saturated hydrocarbon since only single carbon – carbon bonds exist in the molecule.[1]][Polyethyne, however, is an unsaturated hydrocarbon. This is because, although there are single carbon to carbon bonds, there are also carbon to carbon double bonds present.[2]]

- ✓ ✗ I have identified and explained the saturation of polyethene.[1]
- ✓ ✗ I have identified and explained the saturation of polyethyne.[2]

iv [Polyethyne is a conductive polymer as every second bond in its chain is a double bond, which gives rise to delocalised electrons enabling the conduction of electricity.[1]][On the other hand, since polyethene is a saturated hydrocarbon, it does not possess delocalised electrons and hence is not electrically conductive.[2]]

- ✓ ✗ I have explained the impact of the double bond on the electrical conductivity of polyethyne.[1]
- ✓ ✗ I have explained why polyethene does not conduct electricity.[2]

c i [Plasticisers are additives which are found between polymer chains.[1]][By adding plasticisers to a polymer, the ability for chains to pack together is limited with the result that intermolecular bonds are weakened.[2]]

- ✓ ✗ I have identified what plasticisers are and the location in which they are found.[1]
- ✓ ✗ I have identified the impact of plasticisers on the structure of polymers.[2]

ii [The material produced from polymers without plasticisers will be harder and more rigid due to stronger intermolecular bonding.[1]] [On the other hand, the polymer with plasticisers will be softer and more flexible due to weaker intermolecular bonding.[2]]

- ✓ ✗ I have identified the effect of the absence of plasticisers.[1]
- ✓ ✗ I have identified the effect of the presence of plasticisers.[2]

19 a



Amorphous **Crystalline**

b [LDPE is more likely to form amorphous regions.[1]] [Conversely, HDPE is more likely to form crystalline regions.[2]]

- ✓ ✗ I have identified that LDPE is more likely to form amorphous regions.[1]
- ✓ ✗ I have identified that HDPE is more likely to form crystalline regions.[2]

c [LDPE is created at a high temperature and high pressure.[1]] [The highly branched structure of LDPE is produced because the high pressure prevents the formation of neat chains.[2]] [Conversely, HDPE is produced at a high temperature but with lower pressure, which enables polymer chains to be formed in a more orderly fashion, enabling the tight packing of chains.[3]]

- ✓ ✗ I have identified the pressure environment in which LDPE is produced.[1]
- ✓ ✗ I have explained how pressure affects the structure of LDPE.[2]
- ✓ ✗ I have identified the pressure environment in which HDPE is produced and how this affects the structure of HDPE.[3]

d [As LDPE consists primarily of amorphous regions, this type of polymer will allow light to pass through.[1]][This is because LDPE polymer chains are only loosely packed together.[2]][As HDPE consists primarily of crystalline regions, this form of polyethene will appear more opaque than LDPE.[3]][This is because the tight packing of polymer chains in HDPE prevents light being transmitted through it.[4]]

- I have identified the ability of LDPE to transmit light.[1]
- I have described how the structure of LDPE impacts the transmission of light.[2]
- I have identified the inability of HDPE to transmit light.[3]
- I have described how the structure of HDPE impacts the transmission of light.[4]

e [To produce a flexible material, it would be ideal to use LDPE.[1]] [Since polymer chains are less tightly packed in LDPE, the material has greater flexibility compared to HDPE which is harder and more rigid.[2]]

- I have identified that LDPE is the most appropriate polymer to produce a flexible material.[1]
- I have explained why LDPE is more appropriate than HDPE.[2]

Key science skills

20 a [Repeatability refers to the consistency between results obtained using the same process under identical conditions for each trial.[1]] [Process 1 was the least repeatable as there was only consistency between two out of the four results.[2]][Process 2 was the most repeatable process as the result was the same in each trial.[3]] [Process 3 is less repeatable than process 2 but more than process 1 since most trials yielded identical results.[4]]

- I have defined repeatability.[1]
- I have identified the repeatability of process 1.[2]
- I have identified the repeatability of process 2.[3]
- I have identified the repeatability of process 3.[4]

b [Process 2 is most likely carried out at the highest pressure.[1]] [This is because high pressures lead to the formation of more branches in the product which makes the material softer and more flexible.[2]][On the other hand, process 3 is likely carried out at the lowest pressure.[3]][This is because process 3 minimises the amount of branching, enabling polymer chains to pack together tightly to form a hard and brittle material.[4]]

- I have identified the process involving the highest pressure.[1]
- I have explained the impact of high pressures on the properties of the product.[2]
- I have identified the process involving the lowest pressure.[3]
- I have explained the impact of low pressures on the properties of the product.[4]

Chapter 6 review

Multiple choice questions

1 B. Diamond consists of a tetrahedral carbon lattice structure.

2 A. Each carbon atom in graphene is bound to three other atoms producing delocalised electrons.

3 C. Crude oil is a fossil fuel produced over millions of years due to exposure to high pressures.

4 C. Hydrocarbons with triple bonds are alkynes.

5 D. The molecular formula of propane is C_3H_8.

6 A. The parent chain in this molecule is five carbon atoms long.

7 B. Carboxylic acids are not numbered as they are by default on the first carbon atom.

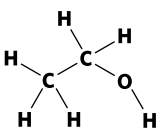
8 A. The component of an ester produced from an alcohol is named first.

9 C. Polymers are repeating units of unsaturated monomers.

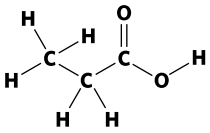
10 B. Crystalline regions are formed in regions which are tightly packed with strong intermolecular bonding.

Short answer questions

11 a



Ethanol



Propanoic acid

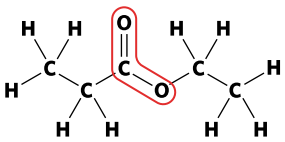
b i [The propanoic acid sample would have a higher boiling point.[1]] [This is because the additional presence of a carbonyl group and a hydroxyl group in the carboxyl group allows for propanoic acid to form, dipole-dipole interactions as well as hydrogen bonds, compared to ethanol being able to only take part in hydrogen bonding.[2]][As a result, propanoic acid would require more energy to convert from a liquid to a gaseous state, leading to a higher boiling point.[3]]

- I have identified that the sample that would be expected to have a higher boiling point.[1]
- I have compared the intermolecular bonds of both compounds.[2]
- I have described the impact of bonding to the boiling point.[3]

ii [The boiling points of ethanol and propanoic acid would both be greater than ethane.[1]][This is because these are both polar molecules, while ethane is non-polar.[2]][This enables the formation of stronger dipole-dipole bonds, which are stronger than the dispersion forces found between ethane molecules.[3]]

- I have compared the boiling points of both compounds in comparison to ethane.[1]
- I have identified the polarity of each molecule.[2]
- I have compared the relative strength of dipole-dipole bonds and dispersion forces.[3]

c i Ethyl propanoate

ii 

iii Sulfuric acid catalyst

iv [The specific name of the reaction is esterification,[1]][and the general name is condensation.[2]]

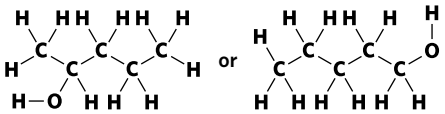
- ✓ ✗ I have identified the specific name of the reaction.[1]
- ✓ ✗ I have identified the general name of the reaction.[2]

d i The boiling point would be lower than both propanoic acid and higher than ethanol.

ii [The addition of water to an ester causes it to split into its alcohol and carboxylic acid components through hydrolysis.[1]][This is because the input of water at high temperatures and in the presence of an acid catalyst causes the covalent bond between the carboxylic acid residue and the alcohol to break, with the atoms in water being used to reform the carboxyl and hydroxyl functional groups back to their original forms.[2]]

- ✓ ✗ I have identified the effect of the addition of water.[1]
- ✓ ✗ I have described the reaction that takes place between an ester and water.[2]

12 a i Butanoic acid and pentan-3-ol

ii 

Pentan-2-ol Pentan-1-ol

b [Thermoplastic polymers can carry out either reprocessing or depolymerisation.[1]][Reprocessing involves shredding, melting and remoulding polymers into new products.[2]][Depolymerisation involves breaking a polymer down into monomers, which are then used to produce new polymers.[3]]

- ✓ ✗ I have identified the types of recycling that thermoplastic polymers can undertake.[1]
- ✓ ✗ I have described the process of reprocessing.[2]
- ✓ ✗ I have described the process of depolymerisation.[3]

13 a i [The structure of diamond is a tetrahedral carbon lattice with each carbon atom bound to four other carbon atoms to form a strong, rigid structure.[1]][Whereas the structure of graphite is a series of layers of carbon atoms each bound to three other carbon atoms in the same layer, with only intermolecular bonding between individual layers.[2]]

- ✓ ✗ I have described the structure of diamond.[1]
- ✓ ✗ I have described the structure of graphite.[2]

ii [Graphite is able to conduct electricity since each carbon atom is bound to three other carbon atoms, leaving one delocalised electron per carbon atom which can pass an electrical current.[1]][Conversely, each carbon atom in diamond is bound to four other carbon atoms, leaving no delocalised electrons and thus acting as an insulator of electricity.[2]][As a result, graphite is most likely to be used in materials that conduct electricity.[3]]

- ✓ ✗ I have explained why graphite is able to conduct electricity.[1]
- ✓ ✗ I have explained why diamond is unable to conduct electricity.[2]
- ✓ ✗ I have linked my answer to the question.[3]

b i Graphite and graphene have the same structure, however graphene has only one layer whereas graphite is multi-layered.

ii Graphene is often referred to as a wonder material due to its exceptional strength, its ability to conduct electricity and its versatility.

c [One commonly seen carbon nanomaterial is nanotubes.[1]][Nanotubes have the shape of a hollow cylinder formed by carbon atoms, and can be used in applications such as antennas due to their ability to conduct electricity, or for strengthening fabrics due to their strength and small size.[2]][Another commonly seen carbon nanomaterial is buckyballs.[3]][Buckyballs have the shape of a hollow sphere formed by carbon atoms, and can be used in applications such as hydrogen storage or anti-cancer agent transport due to their shape.[4]]

- ✓ ✗ I have identified nanotubes as a common nanomaterial.[1]
- ✓ ✗ I have described the shape and some potential uses of nanotubes.[2]
- ✓ ✗ I have identified buckyballs as a common nanomaterial.[3]
- ✓ ✗ I have described the shape and some potential uses of buckyballs.[4]

Key science skills questions

14 a [This experiment is not a reliable measure of the effectiveness of industrial equipment.[1]][A reliable experiment is one that can be reproduced with the same results.[2]][As the results of this experiment's trials vary greatly, this cannot be described as a reliable experiment.[3]]

- ✓ ✗ I have identified that the experiment is not reliable.[1]
- ✓ ✗ I have explained what a reliable experiment is.[2]
- ✓ ✗ I have connected the results of the trials to the question.[3]

b [To determine the success and reliability of an experiment, the same method must be used to avoid variability as well as subjectivity.[1]][For example, using a form of chemical analysis to determine the presence of diamonds would be a better, objective method of determining the presence of diamonds, whereas determining its presence through simple observation will be more subjective.[2]]

- ✓ ✗ I have identified the importance of consistency to determine the results of an experiment.[1]
- ✓ ✗ I have provided examples of objective and subjective forms of chemical analysis.[2]

REVIEW

7A Properties of water

Theory review questions

1 A. Oxygen is a Group 16 element.

2 B. Water has exceptionally high melting and boiling points compared to other Group 16 hydrides due to its ability to form hydrogen bonds.

3 A. Water expands when solid because of the regular hydrogen bonding arrangement that water molecules take up.

4 B. Substances dissolve when the solute-solute and solvent-solvent attractive forces are overcome by the solute-solvent attractive forces.

5 B. Hydrogen bonds are able to form between water molecules and the polar hydroxyl group.

6 B. Ionisation produces charged ions which can then be held in solution by polar water molecules.

7 A. Water, a polar solvent, is often used in biological systems like the human body.

8 B. Solutions provide favourable conditions for chemical reactions to occur.

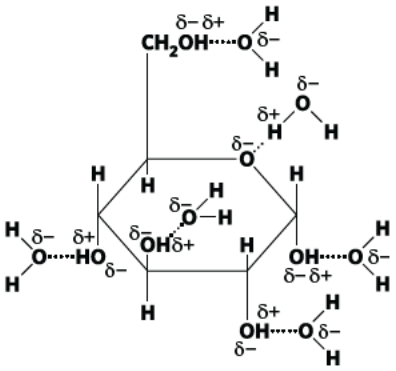
9 B. Substances which are largely non-polar will be unable to dissolve in water.

10 A. All solutions are homogenous mixtures where solute particles are distributed evenly.

Deconstructed questions

11 B 12 C

13



[In order for a substance to dissolve, the solute-solute and solvent-solvent forces need to be broken and the solute-solvent forces need to form.[1]][This is the case, glucose is able to dissolve in water because the solute-solvent forces, hydrogen bonds as illustrated in the diagram, are strong enough to disrupt the solute-solute and solvent-solvent forces.[2]][The chemical equation that represents this process is $C_6H_{12}O_{6(s)} \xrightarrow{H_2O_{(l)}} C_6H_{12}O_{6(aq)}$.[3]]

✓ ✗ I have defined the forces involved when substances dissolve.[1]

✓ ✗ I have identified the solute-solvent forces present and their strength.[2]

✓ ✗ I have included the appropriate chemical equation.[3]

Exam-style questions

Within lesson

14 D 15 D

16 a Group 16 hydrides consist of a Group 16 element like oxygen or sulphur covalently bonded to hydrogen.

b [Molecule A is H_2O, water which has the formula H_2O.[1]][Water has an exceptionally high melting and boiling point, which is representative of the strength of intermolecular forces, compared to other Group 16 hydrides because it is the only Group 16 hydride able to form hydrogen bonds with itself.[2]][This is because the electronegative oxygen atom pulls electrons towards it from the hydrogen atoms, resulting in an electrostatic attraction between the partially positively charged hydrogen atoms and the partially negatively charged oxygen atoms, which results in very strong hydrogen bonds.[3]]

✓ ✗ I have identified the name and formula of molecule A.[1]

✓ ✗ I have identified the intermolecular force responsible for molecule A having high melting and boiling points.[2]

✓ ✗ I have described how this intermolecular force arises.[3]

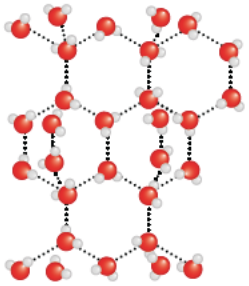
c [The strength of intermolecular forces determines the melting and boiling point of a substance.[1]][Therefore, melting and boiling points increase from Substance B to E as the strength of dispersion forces increases.[2]][The strength of intermolecular dispersion forces increases because as one goes down a group, more electrons are added which makes it more likely for spontaneous dipole-dipole bonds to form.[3]]

✓ ✗ I have explained how melting and boiling points are determined by the strength of intermolecular forces present.[1]

✓ ✗ I have identified the intermolecular force responsible for this increase.[2]

✓ ✗ I have explained how the strength of intermolecular forces increases down a group.[3]

17 a



[When water freezes and becomes ice, the water molecules adopt a crystal-like structure with a regular hydrogen bonding arrangement.[1]][This hydrogen bonding arrangement is shown in the diagram.[2]][Since the water molecules are spaced further apart in this arrangement as opposed to in their liquid state, the volume of ice is greater for the same number of molecules and, therefore, ice has a lower density than water.[3]][Consequently, icebergs are able to float on water as they are less dense than water.[4]]

✓ ✗ I have described the regular hydrogen bonding arrangement in ice.[1]

✓ ✗ I have included a diagram in my response.[2]

✓ ✗ I have described the density and volume of ice compared to water.[3]

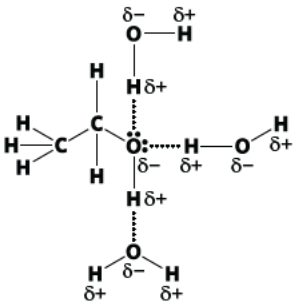
✓ ✗ I have justified why ice floats on water.[4]

b [Since ice floats on the surface of water, it acts as a physical barrier between the freezing, harsh temperatures above the surface of the water and below the surface.[1]][This ensures that marine organisms living under the water have sufficient warmth to survive and do not freeze.[2]]

I have described the role of ice as a physical barrier.[1]

I have explained the significance of this for biological life forms.[2]

c



[Ethanol is soluble in water because it is able to form hydrogen bonds with water molecules as shown in the diagram.[1]]
[Consequently, the solute-solute and solvent-solvent forces can be overcome as new, stronger solute-solvent forces form which include hydrogen bonding interactions.[2]]

I have identified the reason for ethanol's solubility in water.[1]

I have explained the mechanism of dissolving in more detail.[2]

d [Water freezes at 0°C because this is the temperature at which it adopts a stable hydrogen bonding arrangement.[1]][However, since ethanol can also hydrogen bond with water molecules, it can disrupt this stable arrangement, meaning that water will freeze at a colder temperature. Therefore, ethanol can be considered an anti-freezing agent.[2]]

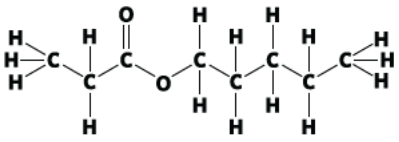
I have explained the process behind the conversion of water to ice.[1]

I have identified how ethanol slows this from happening.[2]

Multiple lessons

18 a Butanoic acid

b

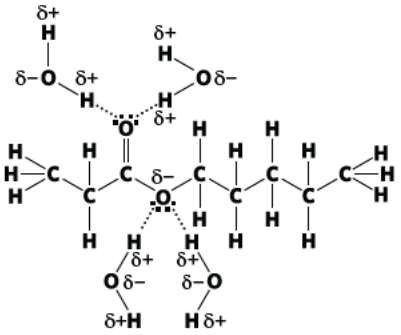


[The ester functional group is present.[1]][The name of this compound is pentyl propanoate.[2]]

I have identified the functional group present.[1]

I have determined the name of this compound.[2]

c



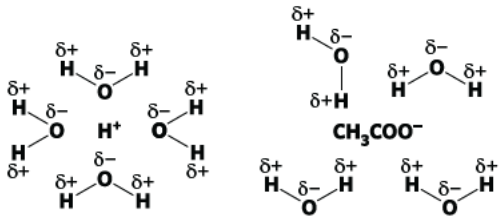
[The ester functional group contains two electronegative oxygen atoms that can therefore form hydrogen bonds with water molecules as shown in the given diagram.[1]][However, while the ester functional group can form hydrogen bonds with water molecules, the vast majority of the molecule is non-polar.[2]][This means that the solute-solvent forces are not enough to disrupt the solute-solute and solvent-solvent forces, meaning that this compound is not soluble in water.[3]]

I have identified the hydrogen bonding between the ester functional group and water molecules with a diagram.[1]

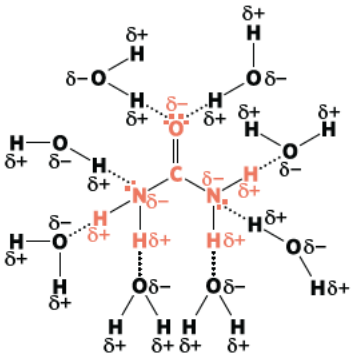
I have identified that the majority of the compound is non-polar.[2]

I have linked this to the solubility of this compound.[3]

d $CH_3COOH_{(l)} \xrightarrow{H_2O_{(l)}} CH_3COO^-_{(aq)} + H^+_{(aq)}$



19 a



[Urea is very soluble in water as it can form hydrogen bonds with water molecules as shown in the diagram.[1]][This means that the solute-solute and solvent-solvent forces can be overcome by the strength of the solute-solvent attractive forces.[2]][This enables urea to dissolve in water as the water molecules 'lift' urea into solution.[3]]

I have described the hydrogen bonding interactions between urea and water molecules with a diagram.[1]

I have explained the mechanism behind the process of a substance dissolving.[2]

I have linked this to the solubility of urea in water.[3]

b [Blood and urine consist mainly of water.[1]][Therefore, since urea is soluble in water, it can freely dissolve into the bloodstream and into urine, allowing it to be transported around the body and then excreted.[2]]

I have recognised the significance of water in the human body.[1]

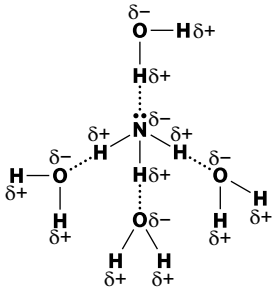
I have described the solubility properties of urea which facilitate its transport.[2]

c [In the process of water reabsorption, water is the solvent and urea is the solute.[1]][This is because the high concentration of urea in the kidney allows water to be reabsorbed from the urine by the process of osmosis.[2]]

I have identified the solvent and solute.[1]

I have justified my answer.[2]

d



[Ammonia, NH_3, is highly soluble in water because it can form hydrogen bonds with water molecules as shown in the given diagram.[1]][Therefore, ammonia is able to dissolve into the bloodstream and be transported to the brain, where it can cause neurological damage.[2]]

I have described the hydrogen bonding interactions between urea and water molecules.[1]

I have explained how its solubility enables it to be transported in the bloodstream.[2]

Key science skills

20 a Qualitative

b Identity of solid compound

c [Validity relates to whether an experimental design accurately tests a hypothesis.[1]][In this case, the experiment is invalid because there is more than one independent variable (both the presence/absence of stirring and the identity of the compound), which means that Margaret cannot compare the solubility differences between compounds.[2]]

I have defined validity.[1]

I have explained why the experiment is invalid.[2]

d Either stir both beakers to the same extent or stir neither beaker to ensure that it remains a controlled variable.

7B Thermal properties of water

Theory review questions

1 A. Different substances have different heat capacities which means they experience different temperature changes given the same amount of heat energy.

2 B. Specific heat capacity is the heat energy required to raise the temperature of 1 g of a substance by 1°C.

3 A. Latent heat is the heat energy required to change the state of 1 mol of a substance.

4 B. Vaporisation is the process where water is converted from a liquid to a gas.

5 A. Fusion is the process where water is converted from a solid to a liquid.

6 B. Latent heat is given the symbol L.

7 A. Specific heat capacity is given the symbol c.

8 A. The latent heat of fusion of water is 6.0 kJ mol^{-1}.

9 B. The latent heat of vaporisation of water is 44.0 kJ mol^{-1}.

10 B. In between points C and D, water is a liquid and so the specific heat capacity of water is used.

Deconstructed questions

11 B **12** D

13 Step 1: melt ice

$q = n \times L$

$n(H_2O) = \frac{150}{(2 \times 1.0) + 16.0} = 8.33$ mol

$L_{fusion}(H_2O) = 6.0$ kJ mol^{-1}

$q = 8.33$ mol $\times$ 6.0 kJ $mol^{-1} = 50$ kJ

Therefore, 50 kJ of heat energy is required to melt 150 g of water.

Step 2: heat water

$q = n \times c \times \Delta T$

The temperature change of water is calculated by $\Delta T = T_{final} - T_{initial}$

$= 100°C - 0°C = 100°C$

The specific heat capacity of water is 4.18 J $g^{-1}°C^{-1}$

$q = 150 \times 4.18 \times 100°C = 62.7 \times 10^3$ J = 62.7 kJ

Therefore, 62.7 kJ of heat energy is required to increase the temperature of water from 0°C to 100°C.

Step 3: boil water

$q = n \times L$

$n(H_2O) = \frac{150}{(2 \times 1.0 + 16.0)} = 8.33$ mol

$L_{vaporisation}(H_2O) = 44.0$ kJ mol^{-1}

$q = 8.33$ mol $\times$ 44.0 kJ $mol^{-1} = 367$ kJ

Therefore, 367 kJ of heat energy is required to boil 150 g of water.

Total energy required = 50 kJ + 62.7 kJ + 367 kJ = 480 kJ

Exam-style questions

Within lesson

14 A **15** B

16 a [Specific heat capacity is the amount of energy required to raise the temperature of 1 gram of a substance by 1°C.[1]][Latent heat is the energy that needs to be absorbed by 1 gram of a substance before a state change can occur (from solid to liquid, or liquid to gas).[2]]

✓ ✗ I have defined specific heat capacity.[1]

✓ ✗ I have defined latent heat.[2]

b [Water molecules are held together by dispersion forces, permanent dipole-dipole forces and hydrogen bonds.[1]][As a result of the strength of these intermolecular forces, predominantly the hydrogen bonds, water absorbs a large amount of energy in its bonds before its temperature increases or before it changes state, resulting in large values for the specific heat capacity and latent heat respectively.[2]]

✓ ✗ I have identified the intermolecular forces present in water.[1]

✓ ✗ I have explained how the intermolecular forces affect specific heat capacity and latent heat.[2]

c [Water's large specific heat capacity enables it to store large amounts of energy in the form of heat before its temperature increases.[1]][This property of water means that excess heat can be removed more efficiently because more energy is transferred to the water molecules per degree of temperature increase.[2]]

✓ ✗ I have identified the property that makes water suitable to remove excess heat.[1]

✓ ✗ I have described how specific heat capacity relates to energy absorbance.[2]

17 a [According to $q = m \times c \times \Delta T$, there is a linear relationship between the temperature change of water and the energy input given a set mass of water.[1]][As a result of this equation and water's specific heat capacity, Walter could have thought that the graph was always a straight line.[2]]

✓ ✗ I have referenced $q = m \times c \times \Delta T$ and the specific heat capacity of water.[1]

✓ ✗ I have linked my answer to the question.[2]

b [One region corresponds to the latent heat of fusion of water (solid to liquid) and has a value of 6.0 kJ mol^{-1}.[1]][The other region corresponds to the latent heat of vaporisation of water (liquid to gas) and has a value of 44.0 kJ mol^{-1}.[2]]

✓ ✗ I have identified one region as the latent heat of fusion of water and given its value.[1]

✓ ✗ I have identified the other region as the latent heat of vaporisation of water and given its value.[2]

c [The latent heat of vaporisation is much greater than the latent heat of fusion.[1]][This is because it takes much more energy to completely disrupt the hydrogen bonds between water molecules to allow them to exist as a gas rather than partially disrupting the hydrogen bonds between water molecules to allow water molecules to exist as a liquid.[2]]

✓ ✗ I have identified which region is larger.[1]

✓ ✗ I have justified my response.[2]

d $n(H_2O) = \frac{25}{(2 \times 1.0) + 16} = 1.39$ mol

$L_{fusion}(H_2O) = 6.0$ kJ mol^{-1}

$q = n \times L = 1.39 \times 6.0 = 8.3$ kJ

Therefore, 8.3 kJ of heat energy is required to melt 25 g of water.

e $n(H_2O) = \frac{25}{(2 \times 1.0) + 16} = 1.39$ mol

$L_{vaporisation}(H_2O) = 44.0$ kJ mol^{-1}

$q = n \times L = 1.39 \times 44.0 = 61$ kJ

Therefore, 61 kJ of heat energy is required to boil 25 g of water.

Multiple lessons

18 a



The hydroxyl group

b [The presence of the polar hydroxyl group in ethanol molecules means that ethanol forms intermolecular dispersion forces, permanent dipole-dipole forces and hydrogen bonds.[1]][Due to the presence of strong hydrogen bonds, ethanol molecules have a high specific heat capacity because these bonds can absorb large amounts of energy.[2]][However, the specific heat capacity of ethanol is less than water because of the presence of the non-polar carbon chain which diminishes the effect of the polar hydroxyl group.[3]]

✓ ✗ I have identified the intermolecular forces present in ethanol.[1]

✓ ✗ I have explained ethanol's relatively high specific heat capacity.[2]

✓ ✗ I have compared the specific heat capacity of ethanol to water.[3]

c $m(H_2O) = 300 \times 0.997 = 299$ g.

Energy released by ethanol = $3 \times 29.6 = 88.8 \times 1000 = 88\,800$ J

$88\,800 = 299 \times 4.18 \times \Delta T$

$\Delta T = \frac{88\,800}{299 \times 4.18} = 71$°C

$\Delta T = T_{final} - T_{initial}$

71°C $= T_{final} - 20$°C

$T_{final} = 91$°C

Therefore, the final temperature of the water is 91°C.

19 **a** [Metallic bonding is the electrostatic attraction between positively charged metal ions and a sea of delocalised electrons.[1]] [Pure metal atoms exist in ordered layers which are surrounded by delocalised electrons in what is referred to as a lattice structure.[2]]

✓ ✗ I have described metallic bonding with respect to ions and electrons.[1]

✓ ✗ I have identified the structure of metals.[2]

b [Heat conduction is the ability for a substance to transfer heat.[1]] [Metals are good heat conductors due to the delocalised electrons and close proximity of metal ions to each other which enables heat, often in the form of kinetic energy, to be rapidly transferred throughout the metal by the electrons moving about or by vibrating metal ions.[2]]

✓ ✗ I have defined heat conduction.[1]

✓ ✗ I have described the properties of metals that enable heat conduction.[2]

c [Specific heat capacity is the amount of energy required to raise the temperature of 1 g of a substance by 1°C.[1]] [Metals have a low specific heat capacity because they conduct heat very effectively which means they require a small amount of heat energy to raise their temperature.[2]]

✓ ✗ I have defined specific heat capacity.[1]

✓ ✗ I have explained why metals have low specific heat capacities.[2]

Key science skills

20 **a** Quantitative

b Heat loss

c Add a lid; bring the beaker of water closer to the flame.

7C Dissolution

Theory review questions

1 B. The majority of ionic compounds can dissolve in water.

2 B. Anions are negatively charged ions, and therefore would interact with the partially positively charged ends of water molecules.

3 B. Ions that dissociate in water become hydrated.

4 B. Dissolved substances contain solute particles (in this case, ions) and solvent particles (in this case, water molecules).

5 A. When ionic substances dissolve in water, they exist as individual ions evenly mixed throughout the solution.

6 B. When sulfate is bonded with Ca^{2+}, Ba^{2+}, Sr^{2+}, Ag^{+}, Hg^{2+} or Pb^{2+} it will be insoluble in water.

7 B. According to the solubility table, PbI_2 is insoluble in water.

8 B. Water is polar and therefore would only be able to dissolve polar substances.

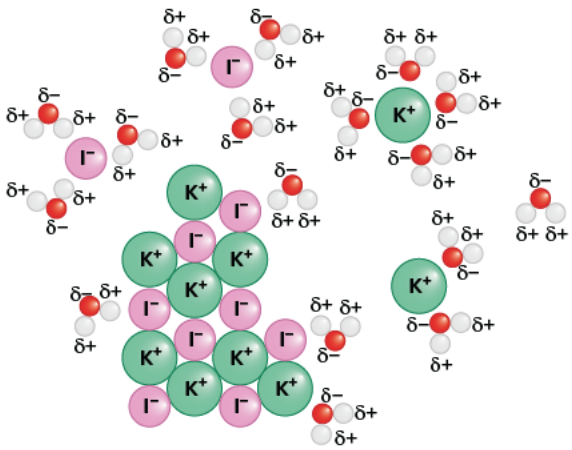
9 A. Since the alcohol is considered soluble, it would be forming hydrogen bonds with water.

10 B. As the size of the non-polar segment of an alcohol increases, the ability of the alcohol to dissolve in water decreases.

Deconstructed questions

11 C **12** C

13



✓ ✗ I have shown the structure of potassium iodide as a solid.

✓ ✗ I have shown the attraction between water molecules and ions at the surface of potassium iodide.

✓ ✗ I have shown individual ions being surrounded by water molecules.

Exam-style questions

Within lesson

14

$HgCl_2$	Na_2CO_3	K_3PO_4	$Ca(OH)_2$
insoluble	soluble	soluble	soluble

15 **a** $MgCl_{2(s)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^{-}_{(aq)}$

b [During dissociation, the negatively charged chloride ions Cl^- at the edges of the ionic lattice become attracted to the partial positively charged ends of the water molecule.[1]] [As a result, an ion-dipole interaction is formed between water and chloride ions.[2]] [Subsequently, chloride ions are pulled away from the lattice structure and become hydrated by being further surrounded by water molecules.[3]]

✓ ✗ I have identified the part of water to which the chloride ions are attracted to.[1]

✓ ✗ I have identified the type of interaction that would occur between water and chloride ions.[2]

✓ ✗ I have described the dissociation of chloride ions from the lattice structure.[3]

16 **a** [This molecule contains a polar hydroxyl group.[1]] [As a result, it is able to form hydrogen bonds with water, providing enough energy to break the intermolecular bonds between the molecules found in the substance and between the water molecules[2]] [and therefore would be able to dissolve in water.[3]]

✓ ✗ I have identified the polar functional group present in the compound.[1]

✓ ✗ I have described the bonding occurring between this molecule and water.[2]

✓ ✗ I have identified the ability of the compound to dissolve in water.[3]

b [This compound contains a polar hydroxyl group, however also contains a long non-polar component.[1]][As a result, the effect of the non-polar chain of the compound will override the polar nature of the polar covalent bond between the O and H.[2]] [Therefore, this compound would not be able to dissolve in water.[3]]

- ✓ ✗ I have described the structure of this compound.[1]
- ✓ ✗ I have described the impact of the structure of the compound on its properties.[2]
- ✓ ✗ I have identified the ability of the compound to dissolve in water.[3]

Multiple lessons

17 a $MgSO_{4(aq)} + 2NaCl_{(aq)} \rightarrow MgCl_{2(aq)} + Na_2SO_{4(aq)}$

b $n(MgSO_4) = \frac{200}{24.3 + 32.1 + (4 \times 16.0)} = 1.66$ mol

c [Magnesium sulfate is an ionic compound, however not all ionic compounds are able to dissolve in water.[1]][Based on the solubility table, the sulfate ion, when bonded with magnesium, is soluble in water.[2]][As a result, epsom salts will be able to dissolve in water.[3]]

- ✓ ✗ I have identified the type of compound represented by $MgSO_4$.[1]
- ✓ ✗ I have identified the solubility of $MgSO_4$.[2]
- ✓ ✗ I have identified the ability of epsom salt to dissolve in water.[3]

d Group 2

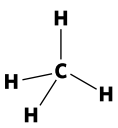
18 a Propanoic acid

b Carboxyl group

c [Propanoic acid contains a carboxyl group, which contains two polar covalent bonds: the O-H bond and the C=O bond.[1]][Both of these components are able to form hydrogen bonds with water. The energy provided by the interaction with water is enough to break the intermolecular bonds between the carboxylic acid as well as the hydrogen bonds between water molecules.[2]][As a result, this molecule would be able to dissolve in water, thereby being soluble in water.[3]]

- ✓ ✗ I have identified the polar covalent bonds in the structure.[1]
- ✓ ✗ I have describe the interactions between propanoic acid and water.[2]
- ✓ ✗ I have identified the solubility of propanoic acid.[3]

19 a



b [Methane is a non-polar hydrocarbon,[1]][therefore is unable to form any strong interactions with polar water molecules.[2]][As a result, it would not be able to dissolve in water.[3]]

- ✓ ✗ I have identified the polarity of methane.[1]
- ✓ ✗ I have identified the intermolecular forces possible between methane and water.[2]
- ✓ ✗ I have identified methane's solubility in water.[3]

c Pentane.

Key science skills

20 a Qualitative data

b Methanoic acid

7D Concentration

Theory review questions

1 A. Solutions with the same concentration have the same solute:solution ratio.

2 B. The units for molarity is mol L^{-1}.

3 B. Solutions with the same concentration have the same solute:solution ratio.

4 B. Ppm shows part per million as shown by mg/L.

5 A. Both values are given in grams, therefore weight would be a better way of expressing concentration.

6 A. Converting the mol to µg gives a 1 : 1 billion ratio (ppb).

7 B. For the values to be compared, the volumes and weights need to be expressed in the same units.

8 A. Dilution occurs when more solvent is added.

9 A. A solution with less solute per same unit of volume is comparatively more dilute.

10 B. Dilution does not affect the amount of solute.

Deconstructed questions

11 A **12** B

13 $n(NaOH) = 2.5 \times \left(\frac{500}{1000}\right) = 1.25$ mol

$m(NaOH) = 1.25 \times (23.0 + 16.0 + 1.0) = 50$ g

Exam-style questions

Within lesson

14 a i $n(CaCl_2) = \frac{1.47}{(40.1 + (2 \times 35.5))} = 0.01323$ mol

$c(CaCl_2) = \frac{0.01323}{\left(\frac{200}{1000}\right)} = 0.0662$ M

ii $n(Na_2CO_3) = \frac{2.59}{(2 \times 23.0) + 12.0 + (3 \times 16.0)} = 0.02443$ mol

$c(Na_2CO_3) = \frac{0.02443}{\left(\frac{250}{1000}\right)} = 0.0977$ M

iii $n(NaCl) = \frac{0.966}{(23.0 + 35.5)} = 0.01651$ mol

$c(NaCl) = \frac{0.01651}{\left(\frac{120}{1000}\right)} = 0.138$ M

iv $c(\text{NaOH}) = \frac{2.0\text{ g}}{200\text{ mL}} = \frac{10\text{ g}}{1000\text{ mL}}$

$m(\text{NaOH}) = 10.0 \times 1000 = 10\,000$ mg

Therefore 10 000 ppm = 1.0×10^4 ppm

v $m(\text{NaOH}) = 2.0 \times (23.0 + 16.0 + 1) = 80.0$ g

$c(\text{NaOH}) = \frac{80.0\text{ g}}{200\text{ mL}} = \frac{400.0\text{ g}}{1000\text{ mL}}$

$m(\text{NaOH}) = 400 \times 1000 = 400\,000$ mg

Therefore 400 000 ppm = 4.0×10^5 ppm

vi $m(FeSO_4) = 0.50 \times (55.8 + 32.1 + (4 \times 16) = 76.0$ g

$\%(w/w) = \frac{76.0}{100} \times 100 = 76\%$

b i $1.0\text{ M} = \frac{1\text{ mol}}{1\text{ L}}$

$m(\text{NaCl}) = 1.0 \times (23.0 + 35.5) = 58.5 \times 1000 = 58\,500$ mg

Therefore 5.9×10^4 ppm

ii $1337 \times 1000 = 1.337 \times 10^6$ ppb

iii $981\text{ ppm} = \frac{981\text{ mg}}{1\text{L}}$

$n(\text{KCl}) = \frac{\left(\frac{981}{1000}\right)}{(39.1 + 35.5)} = 0.01315$ mol

$c(\text{KCl}) = \frac{0.01315}{1\text{ L}} = 0.0132$ M

iv $1.5\text{ M} = \frac{1.5\text{ mol}}{1\text{ L}}$

$m(PbI_2) = 1.5 \times (207.2 + (2 \times 126.9) = 691.5$ g

$\%(m/v) = \frac{691.5\text{ g}}{(1 \times 1000)} \times 100 = 69\%$

c i $1.25 \times \left(\frac{20.0}{1000}\right) = c_2 \times \left(\frac{25.0 + 20.0}{1000}\right)$

$c_2 = 0.556$ M

ii $c_1 \times \left(\frac{20.0}{1000}\right) = 2.0 \times \left(\frac{50.0}{1000}\right)$

$c_1 = 5.0$ M

iii $1.05 \times \left(\frac{150}{1000}\right) = c_2 \times \left(\frac{200 + 150}{1000}\right)$

$c_2 = 0.450$ M

15 a 8.7 g in every 100 mL

$\%(m/v) = \frac{8.7}{100} \times 100 = 8.7\ \%$

b $n(\text{sugar}) = \frac{8.7}{180} = 0.0483$ mol

$c(\text{sugar}) = \frac{0.0483}{\left(\frac{100}{1000}\right)} = 0.48$ M

16 a 1 ppm = 1 mg/kg = 0.5 mg/500 g

Therefore in a 500 g fish, the maximum level of mercury that can be found is 0.5 mg of mercury.

b 1.76 M = 1.76 moles of mercury for every litre.

$m(\text{Hg}) = 1.76 \times 200.6 = 353.1$ g

Therefore in 50 mL:

$m(\text{Hg}) = 353.1 \times \left(\frac{50.0}{1000}\right) = 17.7$ g

c 17.7 g = 17.7 × 1000 = 17 700 mg of mercury

$\text{ppm} = \frac{17\,700}{2.67} = 6629.2 = 6.63 \times 10^3$ ppm

[The fish has a mercury level of 6.63×10^3 ppm,[1]][which is much higher than the Australian requirement of 1 ppm.[2]][Therefore this fish would not meet the Australian standards.[3]]

✓ ✗ I have included calculations in my answer.

✓ ✗ I have identified the concentration (in ppm) of mercury in the fish.[1]

✓ ✗ I have compared the ppm in the fish with Australian standards.[2]

✓ ✗ I have linked my answer to the question.[3]

d $c_1 = 1.76$ M

$V_1 = \left(\frac{50.0}{1000}\right) = 0.050$ L

$c_2 = 1.0$ M

$V_2 = ?$

$V_2 = \frac{1.76 \times 0.050}{1.0} = 0.088$ L

The volume required = 0.088 − 0.050 = 0.038 L

17 a 50 %m/v = 50 g/100 mL

$m_{\text{per 1L}} = 50 \times 10 = 500$ g

b $n(C_6H_{12}O_6) = \frac{500}{(6 \times 12.0) + (12 \times 1.0) + (6 \times 16.0)} = 2.78$ mol

$c = \frac{2.78}{1.0\text{ L}} = 2.78$ M

c $c_1 = 2.78$ M

$V_1 = \frac{50.0}{1000} = 0.050$ L

$c_2 = ?$

$V_2 = \frac{200}{1000} = 0.200$ L

$c_2 = \frac{2.78 \times 0.050}{0.200} = 0.70$ M

18 a 7.0 ppb = 7.0 µg per 1kg

Mass = 7.0 × 13.6 = 95 µg

b $n(\text{aldicarb})_{200\text{ mL}} = 1.82 \times 10^{-5} \times \left(\frac{200}{1000}\right) = 3.64 \times 10^{-6}$ mol per 200 mL

$m(\text{aldicarb}) = 3.64 \times 10^{-6} \times 190.3 = 6.926 \times 10^{-4}$ g

$m(\text{aldicarb}) = 6.926 \times 10^{-4} \times 1\,000\,000 = 693$ µg

[By consuming 200mL of the 1.82×10^{-5} M aldicarb water sample, the toddler would have consumed a total of 693 µg[1]][which is more than the 95.2 µg recommended by the guidelines.[2]]

✓ ✗ I have included calculations in my answer.

✓ ✗ I have identified the total aldicarb consumed.[1]

✓ ✗ I have linked my answer to the question.[2]

Multiple lessons

19 a Ionic compound

b $KCl_{(s)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)}$

c $n(KCl) = 2.5 \times \frac{100}{1000} = 0.25$ mol

$m(KCl) = 0.25 \times (39.1 + 35.5) = 18.65 = 18.7$ g

d $AgNO_{3(aq)} + KCl_{(aq)} \rightarrow AgCl_{(s)} + KNO_{3(aq)}$

Key science skills

20 a Personal error

b [Since there was some mass of NaCl lost during the experiment, there would be less NaCl than required.[1]][Therefore the actual concentration of the solution would be less than the expected value.[2]]

✓ ✗ I have identified the effect of the error on the sample.[1]

✓ ✗ I have identified the effect of the error on the results.[2]

c $n(NaCl) = \frac{4.00}{(23.0 + 35.5)} = 0.06838$ mol

$c(NaCl) = \frac{0.06838}{\left(\frac{250}{1000}\right)} = 0.274$ M

7E Solubility

Theory review questions

1 B. The solubility of gaseous solutes and some solid solutes decreases as the temperature increases.

2 B. A saturated solution is a solution unable to dissolve any more solute.

3 A. Crystallisation takes place when a supersaturated solution is disturbed or when a saturated solution is cooled.

4 B. Solubility tables are used to predict the solubility of ionic compounds.

5 A. Solubility curves are constructed to visualise the relationship between temperature and solubility as these are rarely linear.

6 A. Unsaturated solutions are those that do not have enough solute dissolved to be saturated.

7 B. Increasing the temperature tends to decrease the solubility of gases.

8 B. Solubility is the extent to which a solute is able to dissolve in a solvent.

9 B. The line represents a saturated solution, and above this line will be supersaturated solutions.

10 A. Solubility curves show the maximum amount of solute that can be dissolved in a given quantity of solvent at a particular temperature.

Deconstructed questions

11 C **12** A

13 [At 90°C, the solubility of $K_2Cr_2O_7$ is 70 g per 100 g of H_2O. As there are 200 g of H_2O in the solution, this means that 140 g of $K_2Cr_2O_7$ will be found in a saturated solution.[1]][In a supersaturated solution, the amount of solute dissolved is greater than the amount found in a saturated solution for a given amount of solvent.[2]][As there is 180 g of $K_2Cr_2O_7$ dissolved in the 200 g of solvent, and the saturated amount is 140 g, there will be 40 g of crystals formed when the solution is disturbed.[3]]

✓ ✗ I have identified the amount of $K_2Cr_2O_7$ present in a saturated solution.[1]

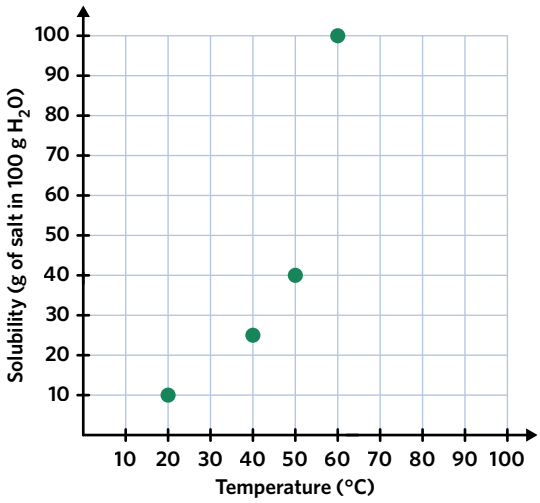
✓ ✗ I have defined a supersaturated solution.[2]

✓ ✗ I have identified the mass of crystals formed from this solution.[3]

Exam-style questions

Within lesson

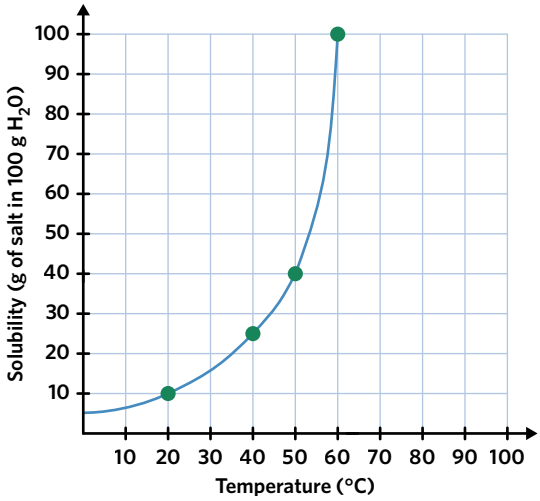
14 a



✓ ✗ I have labelled my axes.

✓ ✗ I have plotted the values on the graph.

b



✓ ✗ I have drawn the solubility curve by connecting the data points.

c [The solubility curve has the shape of an exponential curve.[1]] [This implies that as the temperature increases, the solubility of the solute will increase exponentially.[2]]

✓ ✗ I have commented on the shape of the solubility curve.[1]

✓ ✗ I have described the implication of the solubility curve's shape.[2]

15 a i Soluble

ii Insoluble

iii Insoluble

iv Soluble

b [Emily's claim is incorrect.[1]][The formation of solids as solute is continuously added to a solution does not necessarily indicate that the solute is insoluble in the solvent, instead it may be an indication of supersaturation.[2]][As a supersaturated solution is unstable, the excess solute added to the solution will crystalise when disturbed, forming solids.[3]]

- ✓ ✗ I have identified the accuracy of Emily's claim.[1]
- ✓ ✗ I have explained how solids can be found as a result of supersaturation.[2]
- ✓ ✗ I have identified the effect of supersaturation.[3]

c [If an insoluble compound were added to water, it would not disperse throughout the water, instead remaining bound to itself.[1]] [The main difference between a solid from an insoluble compound and a solid produced through crystallisation is that crystals, despite not being dissolved, are still made of a soluble compound while the insoluble compound is not.[2]][Additionally, an insoluble compound always remains in the solid state and does not crystallise.[3]] [Crystallisation, however, only occurs once the solution is supersaturated with a soluble compound.[4]]

- ✓ ✗ I have described what would happen if an insoluble compound were added to water.[1]
- ✓ ✗ I have identified the chemical differences between the insoluble compound and the crystals.[2]
- ✓ ✗ I have identified the relationship between insoluble compounds and crystals.[3]
- ✓ ✗ I have described the formation of crystals.[4]

16 a The most common relationship between temperature and solubility is that as the temperature increases, the solubility of a solute also increases.

b [The ionic compound $Ce_2(SO_4)_3$ does not follow the most common relationship between temperature and solubility in ionic compounds.[1]] [This is because the solubility of $Ce_2(SO_4)_3$ decreases as the temperature increases initially, and then remains stable.[2]]

- ✓ ✗ I have identified that $Ce_2(SO_4)_3$ is an outlier on the solubility curve.[1]
- ✓ ✗ I have described the solubility curve of $Ce_2(SO_4)_3$.[2]

c i Unsaturated

ii Saturated

iii Saturated

iv Supersaturated

v Supersaturated

17 a [Firstly, the student should weigh the solute so that each test tube has the same mass of solute being introduced.[1]][The student should then ensure that the volume of water in each test tube is the same.[2]] [Following this, the water bath should be heated while stirring the contents of the test tubes until the solutes have entirely dissolved, and then allowed to cool down until crystals begin to form.[3]] [The temperature at which this occurs should be recorded, showing the point of saturation.[4]][This experimental technique should be repeated to ensure validity of results.[5]]

- ✓ ✗ I have identified that the mass of solute weighed should be kept consistent across test tubes.[1]
- ✓ ✗ I have explained that the volume of solvent should be consistent.[2]
- ✓ ✗ I have described the importance of constantly stirring the test tubes and heating them in a water bath.[3]
- ✓ ✗ I have described that the temperature should be taken at the point of crystallisation.[4]
- ✓ ✗ I have identified why the experiment should be repeated.[5]

b i Unsaturated

ii Supersaturated

iii Cannot be determined

Multiple lessons

18 a i A solute is a substance that is dissolved in a solution.

ii A solvent is a substance in which a solute is dissolved.

b [Ethanol will have a higher solubility in water than ethane.[1]][This is because ethanol possesses a hydroxyl group, -OH, which makes the molecule more polar and enables it to form strong intermolecular hydrogen bonds with water, which can overcome the forces between water molecules, thus enabling dissolution in water.[2]] [In comparison, ethane is non-polar and thus only able to form weaker dispersion forces. Consequently, ethane is unable to overcome the intermolecular forces between water molecules, making it less soluble in water.[3]]

- ✓ ✗ I have compared the solubility of ethanol and ethane in water.[1]
- ✓ ✗ I have described the intermolecular bonding of ethanol and how it relates to dissolution.[2]
- ✓ ✗ I have described the intermolecular bonding of ethane and how it relates to dissolution.[3]

c i [The solubility curve for ethane does have an expected relationship between temperature and solubility.[1]][As ethane has a boiling point of −89°C, the molecule will be in a gaseous state for all the temperatures shown on the solubility curve.[2]][As the general trend for gaseous molecules on the solubility curve is that they decrease in solubility as the temperature increases, this means that the solubility curve of ethane has a conventional shape.[3]]

- ✓ ✗ I have identified the trend of the solubility curve of ethane.[1]
- ✓ ✗ I have described the state of ethane based on the conditions of the solubility curve.[2]
- ✓ ✗ I have described the common relationship between temperature and solubility for gaseous molecules.[3]

ii [As ethanol is not in a gaseous form throughout much of the solubility curve, the solubility would be expected to increase as the temperature increases.[1]][This is the opposite of ethane, which decreases in solubility as the temperature increases.[2]] [However, this trend is not universal, and some non-gaseous molecules that are soluble in water do not increase in solubility as temperature increases, so this cannot be predicted with certainty.[3]]

I have identified the solubility of ethanol as temperature increases.[1]

I have compared the expected solubility curve of ethanol to that of ethane.[2]

I have identified whether the solubility curve of ethanol can be predicted with certainty.[3]

d i According to the graph, 30 g $K_2Cr_2O_7$ dissolves in 100 g H_2O at 50°C.

$\frac{500}{100} = 5$

$30 \times 5 = 150$ g $K_2Cr_2O_7$ dissolved in 500 g H_2O at 50°C.

ii $n(K_2Cr_2O_7) = \frac{150}{(2 \times 39.1) + (2 \times 52.0) + (7 \times 16.0)} = 0.510$ mol

iii 500g of water = 500 mL

$c(K_2Cr_2O_7) = \frac{0.51}{0.500} = 1.0$ mol L^{-1}

19 a [This is not an effective method of producing a solubility table.[1]] [The student leaving the solution unobserved for ten minutes is an uncontrolled variable, as it is not possible to account for any interactions during this time.[2]] [Additionally, some solutes may dissolve more quickly than others, and therefore by leaving a solution for only ten minutes, the experimenter may determine that it has a lower solubility, when there may simply have not been enough time for complete dissolution.[3]] [Lastly, the method of Riley's observation is highly subjective, and this may produce uncontrolled variability in the observations recorded.[4]]

I have identified that this is not the most effective method of producing a solubility table.[1]

I have explained that the student not observing the solution is an uncontrolled variable.[2]

I have explained that the experiment may produce inaccuracies on observed relative solubility.[3]

I have identified that Riley's observation method is subjective.[4]

b i H – Hydrogen, Li – Lithium, Na – Sodium, K – Potassium, Rb – Rubidium, Cs – Caesium, Fr – Francium

ii All group IA elements are soluble in water.

Key science skills

20 a A random error.

b Examples include:

- Using the wrong ionic compound.
- Incorrectly measuring the solubility of a compound, such as incorrectly determining the point of saturation.
- Variations in temperature.
- Incorrectly measuring the mass of the solute.

c [As the results are all centred around the predicted result, the results are accurate.[1]] [However, as the results are all different to each other, they are not precise.[2]]

I have commented on the accuracy of the results.[1]

I have commented on the precision of the results.[2]

d [As the results of his experiment are imprecise, the experimental results can be improved through refining the experimental technique.[1]] [For example, precision may be improved by cooling the water more slowly and ensuring that stirring is taking place,to remove a degree of variability in the results, as well as using better apparatus and calculating results to more significant figures.[2]]

I have identified how to improve the experimental results.[1]

I have provided an example of how to achieve this improvement.[2]

e [Stirring the solution during heating and cooling ensures consistency in the distribution of molecules and the chemical interactions in the solution.[1]] [This consistency reduces the likelihood of random variability between the results, thus reducing imprecision as all of the solutions will be expected to have more similar chemical interactions.[2]]

I have described the effect of stirring on a solution.[1]

I have explained why consistency is important.[2]

7F Precipitation reactions

Theory review questions

1 B. Precipitation reactions produce solids from aqueous solutions.

2 B. The formation of a solid precipitate turns the solution cloudy.

3 A. A precipitation reaction can be identified by the formation of a solid from aqueous solutions.

4 A. Barium sulphate is insoluble in water.

5 A. Full equations include all chemical species present.

6 B. Ionic equations represent only the chemical species reacting to produce the precipitate.

7 B. Spectator ions do not react and remain in the aqueous state on both sides of the equation.

8 A. The solid produced in a precipitation reaction can be easily removed from solution.

9 A. AgCl is insoluble in water.

10 B. Ionic equations represent only the species reacting to produce the precipitate.

Deconstructed questions

11 A **12** C

13 Full equation: $Co(NO_3)_{2(aq)} + 2NaOH_{(aq)} \rightarrow Co(OH)_{2(s)} + 2NaNO_{3(aq)}$

Ionic equation: $Co^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Co(OH)_{2(s)}$

Spectator ions: $Na^+_{(aq)}$, $NO_3^-{}_{(aq)}$

Exam-style questions

Within lesson

14 a The industrial chemist should observe the test tube becoming cloudy as a solid precipitate is formed.

b [The visual change is evidence of a precipitation reaction as a solid is formed from the colourless, aqueous solutions.[1]][The potential ionic compounds that could be formed when $Ca^{2+}_{(aq)}$ is added to each test tube are $CaCl_2$, $Ca(NO_3)_2$ and $CaSO_4$.[2]][According to the solubility rules, $CaCl_2$ and $Ca(NO_3)_2$ are both soluble in water while $CaSO_4$ is insoluble.[3]][Since $CaSO_4$ is insoluble in water, a precipitation reaction will occur and therefore, solution B contains $SO_4^{2-}{}_{(aq)}$ ions.[4]]

- I have linked the observation to a specific type of chemical reaction taking place.[1]
- I have identified the possible ionic compounds that could form.[2]
- I have determined the solubility of each of these ionic compounds.[3]
- I have identified the ion present in solution B.[4]

c $Ca^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow CaSO_{4(s)}$

d [The visual change is evidence of a precipitation reaction as a solid is formed from the colourless, aqueous solutions.[1]][Since solution B contains $SO_4^{2-}{}_{(aq)}$ ions, solution A must either contain $NO_3^-{}_{(aq)}$ or $Cl^-_{(aq)}$ ions. Therefore, the possible ionic compounds that could form are $AgNO_3$ or AgCl.[2]][According to the solubility rules, AgCl is insoluble while $AgNO_3$ is soluble in water.[3]][Since a precipitation reaction occurs in solution A and not solution C, solution A must contain $Cl^-_{(aq)}$ ions and solution C must contain $NO_3^-{}_{(aq)}$ ions.[4]]

- I have linked the observation to a specific type of chemical reaction taking place.[1]
- I have identified the possible ionic compounds that could form.[2]
- I have determined the solubility of each of these ionic compounds.[3]
- I have identified the ions that are present in solution A and C.[4]

e $Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$

15 a $Ca^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow CaCO_{3(s)}$

b [Pipe blockages can be caused by the formation of limescale, $CaCO_{3(s)}$ precipitate.[1]][Since areas with higher concentrations of $Ca^{2+}_{(aq)}$ will have a greater rate of formation of $CaCO_{3(s)}$, the accumulation of limescale will lead to pipe blockages occurring more frequently.[2]]

- I have identified one of the causes of blocked pipes.[1]
- I have linked the increased calcium levels to increased limescale accumulation.[2]

c [The ion exchange resin replaces $Ca^{2+}_{(aq)}$ ions with $Na^+_{(aq)}$ ions in water which reduces the concentration of $Ca^{2+}_{(aq)}$ and increases the concentration of $Na^+_{(aq)}$.[1]][Since limescale is caused by the formation of $CaCO_{3(s)}$, a decrease in $Ca^{2+}_{(aq)}$ concentration will reduce the formation of limescale as $Ca^{2+}_{(aq)}$ ions are replaced with $Na^+_{(aq)}$, forming the highly soluble compound $Na_2CO_{3(aq)}$ instead.[2]]

- I have identified the role of the ion exchange resin.[1]
- I have explained the function of the ion exchange resin to the reduction of limescale accumulation.[2]

d [The reaction of limescale with a strong acid results in the breakdown of a solid and the formation of a soluble salt ($CaCl_{2(aq)}$), water ($H_2O_{(l)}$) and carbon dioxide gas ($CO_{2(g)}$).[1]][Since none of the products are solids, there will be a reduction in pipe blockages caused by limescale as the solid is reacted to produce products that can be easily flushed away in pipes.[2]]

- I have identified the products of the reaction and their states.[1]
- I have explained the effect of the reaction on pipe blockages.[2]

16 a $Ca^{2+}_{(aq)} + C_2O_4^{2-}{}_{(aq)} \rightarrow CaC_2O_{4(s)}$

b [Dehydration means that there is less water present in the kidney, increasing the concentration of $Ca^{2+}_{(aq)}$ and $C_2O_4^{2-}{}_{(aq)}$, which results in greater levels of $CaC_2O_{4(s)}$.[1]][Furthermore, since compounds are less likely to dissolve with decreased water levels, more $CaC_2O_{4(s)}$ will form and consequently, more kidney stones.[2]]

- I have described the effects of decreased water levels on $CaC_2O_{4(s)}$ formation.[1]
- I have linked the formation of $CaC_2O_{4(s)}$ precipitate to the development of kidney stones.[2]

c [$Ca^{2+}_{(aq)}$ ions are one of the reactants required for the formation of $CaC_2O_{4(s)}$ precipitate.[1]][Therefore, high levels of calcium in the kidney can increase the amount of $CaC_2O_{4(s)}$ formed, leading to an increased risk of developing kidney stones.[2]]

- I have described $Ca^{2+}_{(aq)}$ as a necessary reactant for precipitate formation.[1]
- I have linked increased precipitation formation to an increased risk of kidney stones.[2]

d [The presence of the other ions mentioned will not affect the ability for precipitation reactions to be used.[1]][This is because the other ions form highly soluble ionic compounds (as opposed to $Ca^{2+}_{(aq)}$) and so will not precipitate out of the solution and interfere with the results.[2]]

- I have determined the effect of other ions on the formation of a precipitate.[1]
- I have justified my response with reference to the solubility of the ionic compounds formed.[2]

e Full equation:

$Ca^{2+}_{(aq)} + Na^+_{(aq)} + K^+_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow CaCO_{3(s)} + Na^+_{(aq)} + K^+_{(aq)}$

17 a Full equation: $BaS_{(aq)} + H_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + H_2S_{(aq)}$

Spectator ions: $S^{2-}_{(aq)}$ and $H^+_{(aq)}$

b Ionic equation: $Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$

c [Barium sulfate is safe for human consumption because it is a solid, as opposed to other soluble compounds of barium.[1]] [This means that the highly toxic barium is locked in a solid and so can pass through the human digestive system without being absorbed by the body.[2]]

- ✓ ✗ I have identified the insoluble nature of barium sulfate.[1]
- ✓ ✗ I have linked barium sulfate's insolubility to it not being absorbed by the body.[2]

d [Solutions with precipitate are often cloudy and since barium sulfate is a white solid, it will have the appearance of milk.[1]] [Since barium sulfate is insoluble, it will exist in a solid form, which will make the solution very dense like a 'thick glass of milk' due to the undissolved solute.[2]]

- ✓ ✗ I have described the appearance of solutions with a precipitate.[1]
- ✓ ✗ I have described the density of solutions with a precipitate.[2]

Multiple lessons

18 a $A_{(aq)} + B_{(aq)} \rightarrow C_{(s)}$

b [A precipitation reaction is a chemical reaction where a solid forms from an aqueous solution.[1]] [In this case, since the antigen and antibody are both in the aqueous state, and when they react, they produce a solid antigen-antibody complex, this is classified as a precipitation reaction.[2]]

- ✓ ✗ I have defined a precipitation reaction.[1]
- ✓ ✗ I have linked the definition of a precipitation reaction to the reaction occurring.[2]

c [Patient A has COVID-19.[1]] [This is because the precipitation reaction that occurred, evidenced by the formation of a solid, shows that the antigen for COVID-19 was present and reacted with its respective antibody.[2]]

- ✓ ✗ I have identified which patient has COVID-19.[1]
- ✓ ✗ I have justified my answer with reference to precipitation reactions.[2]

d [Solubility increases as temperature increases because the solvent particles can more effectively break apart the solute particles.[1]] [Therefore, since diagnosis is based upon the formation of a precipitate, this would be less easily observed at higher temperatures because solubility increases, meaning less solid precipitate would form.[2]]

- ✓ ✗ I have explained the effect of temperature on solubility.[1]
- ✓ ✗ I have justified why cold temperatures are preferred for rapid diagnosis tests.[2]

19 a Full equation: $Ca(OH)_{2(aq)} + Mg^{2+}_{(aq)} \rightarrow Mg(OH)_{2(s)} + Ca^{2+}_{(aq)}$

Ionic equation: $Mg^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Mg(OH)_{2(s)}$

b Full equation: $Na_2CO_{3(aq)} + Ca^{2+}_{(aq)} \rightarrow CaCO_{3(s)} + 2Na^{+}_{(aq)}$

Ionic equation: $Ca^{2+}_{(aq)} + CO_3{}^{2-}_{(aq)} \rightarrow CaCO_{3(s)}$

c [The removal of $Mg^{2+}_{(aq)}$ ions should be performed first.[1]] [This is because since $Ca^{2+}_{(aq)}$ is introduced to the solution with the addition of lime, these ions have to be removed. This can occur in the next step, by the addition of soda ash to produce insoluble calcium carbonate.[2]]

- ✓ ✗ I have identified the precipitation reaction that should be carried out first.[1]
- ✓ ✗ I have justified my answer with reference to the ions added by each chemical compound.[2]

d $M(Na_2CO_3) = (2 \times 23.0) + 12.0 + (3 \times 16.0) = 106.0 \text{ g mol}^{-1}$

$n(Na_2CO_3) = \frac{5.0}{106.0} = 0.047 \text{ mol}$

e $n(CaCO_3) = n(Na_2CO_{3(aq)}) = 0.047 \text{ mol}$

$M(CaCO_3) = 40.0 + 12.0 + (3 \times 16.0) = 100.0 \text{ g mol}^{-1}$

$m(CaCO_3) = 0.047 \times 100.0 = 4.7 \text{ g}$

Key science skills

20 a Qualitative

b [Ionic compounds containing either cobalt or manganese would both appear pink according to the table.[1]] [Consequently, this would introduce a systematic error as it is impossible, with the current experimental design, to distinguish between the two metal cations.[2]]

- ✓ ✗ I have identified the similarities between the two metals based on their colour.[1]
- ✓ ✗ I have described the type of error present.[2]

c [Precision is a measure of how close values are to each other.[1]] [Since concentration based on cloudiness is subjective, the values determined would not be very precise.[2]] [Reliability is a measure of the reproducibility of results.[3]] [Since concentration based on cloudiness depends on the experimenter who is recording the data, it would be hard to get reliable results each time the experiment is repeated.[4]] [Accuracy is a measure of how close the experimental values are to the true value.[5]] [Since concentration based on cloudiness is purely qualitative rather than quantitative, the results generated would not be accurate and would only give an indication of the relative concentrations of metals, rather than precise numerical values.[6]]

- ✓ ✗ I have defined precision.[1]
- ✓ ✗ I have identified whether the experimental technique is precise or not.[2]
- ✓ ✗ I have defined reliability.[3]
- ✓ ✗ I have identified whether the experimental technique is reliable or not.[4]
- ✓ ✗ I have defined accuracy.[5]
- ✓ ✗ I have identified whether the experimental technique is accurate or not.[6]

Chapter 7 review

Multiple choice questions

1 C. All Group 16 hydrides have a bent shape according to VSEPR.

2 C. Boiling from 20°C requires the specific heat capacity of water as well as its latent heat of vaporisation.

3 D. Solvents are substances which dissolve another substance.

4 C. Molarity expresses concentration in mol/L.

5 D. Supersaturated solutions have more solute dissolved than what the solubility curve will say.

6 A. Silver chloride is an ionic compound that is insoluble in water.

7 D. %(w/v) is a unit of concentration equivalent to g/100 mL.

8 B. Precipitation reactions result in the formation of a solid from an aqueous solution.

9 A. Stirring a solution does not change the solubility of a solute but instead only increases the rate at which the solute dissolves.

10 C. Solubility curves are generated by recording the point at which crystals start to form when cooling a solution down.

Short answer questions

11 a [Water's unique properties compared to other Group 16 hydrides are due to its hydrogen bonding capabilities.[1]][It has an exceptionally high melting and boiling point because hydrogen bonds are a very strong form of intermolecular bonding, which are absent in other Group 16 hydrides. This means that more energy is required to disrupt the intermolecular forces of attraction between water molecules, leading to a greater melting and boiling point.[2]][Water also increases in volume (decreases in density) as it freezes. This is because the water molecules adopt a regular, crystal-like lattice structure, spaced out further apart and hydrogen bonded to four other water molecules.[3]]

- I have identified water's uniqueness as being due to hydrogen bonding.[1]
- I have referenced and explained water's high melting and boiling point.[2]
- I have referenced and explained ice's decreased density compared to liquid water.[3]

b $q = n \times L$

$100\text{ kJ} = n\text{ mol} \times 44.0\text{ kJ mol}^{-1}$

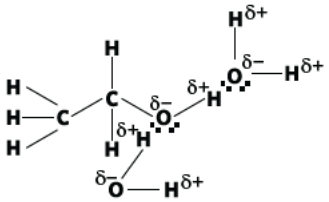
$n(H_2O) = \frac{100}{44.0} = 2.27\text{ mol}$

$m(H_2O) = 2.27 \times ((2 \times 1.0) + 16.0) = 40.9\text{ g}$

$V(H_2O) = \frac{40.9}{0.997} = 41.0\text{ mL}$

Therefore, 41.0 mL of water must be perspired.

c [$CH_3CH_2OH_{(l)} \xrightarrow{H_2O_{(l)}} CH_3CH_2OH_{(aq)}$.[1]][Ethanol is able to dissolve in water because the powerful hydrogen bonding interactions, as shown in the diagram, are able to overcome the weaker intermolecular forces of attraction between the solute molecules and between the solvent molecules.[2]]



- I have given the equation for the dissolution of ethanol.[1]
- I have explained why ethanol is able to dissolve in water.[2]
- I have drawn the hydrogen bonding interactions possible between ethanol and water.

d $M(C_2H_5OH) = (2 \times 12.0) + (6 \times 1.0) + 16.0 = 46.0\text{ g mol}^{-1}$

$n(C_2H_5OH) = \frac{0.05}{46.0} = 0.001087\text{ mol}$

$c(C_2H_5OH) = \frac{0.001087}{\left(\frac{100}{1000}\right)} = 0.01\text{ M}$

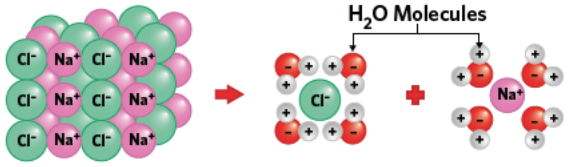
e [Oxygen is a non-polar molecule.[1]][Therefore, the only solute-solvent attractive forces are dispersion forces. The strength of the dispersion forces between oxygen molecules and water molecules is therefore not strong enough to overcome the strong hydrogen bonding between water molecules and, therefore, oxygen is transported in the blood by proteins such as hemoglobin instead.[2]]

- I have described the polarity of oxygen.[1]
- I have explained why oxygen requires hemoglobin to be transported around the body.[2]

f [The high specific heat capacity of water means that it takes a large amount of energy to raise the temperature of 1 g of water by 1°C, meaning that the temperature of the ocean can remain fairly constant as a lot of energy is required to be absorbed before any significant temperature change occurs.[1]][The ability for ice to float on water due to its lower density than liquid water also protects the water underneath from the freezing temperatures above, acting as a protective layer to stop the ocean from completely freezing in some parts of the world.[2]]

- I have described the high specific heat capacity of water and explained how this influences water temperature.[1]
- I have described ice's ability to float on water and explained how this influences water temperature.[2]

12 a [$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$.[1]][NaCl is able to dissolve in water because the polar H_2O molecules are able to pull apart the NaCl crystal lattice to create a solution of dissolved sodium and chloride ions each surrounded by a shell of water molecules.[2]]



- I have written the equation for the dissociation of NaCl.[1]
- I have explained the dissociation of NaCl.[2]
- I have included a diagram in my response.

b A supersaturated solution is a solution that contains more than the normal amount of solute dissolved at a given temperature.

c [A concentration greater than approximately 37.8 g of NaCl/100 mL of water is required to create a supersaturated solution at 80°C.[1]] [When the solution is disturbed, crystallisation may occur.[2]]

✓ ✗ I have used the solubility curve to determine the concentration of NaCl required to generate a supersaturated solution.[1]

✓ ✗ I have identified what may happen when the solution is disturbed.[2]

d Concentration required = 37 g NaCl/100 mL of water

$M(\text{NaCl}) = 23.0 + 35.5 = 58.5 \text{ g mol}^{-1}$

$n(\text{NaCl}) = \frac{37}{58.5} = 0.6325 \text{ mol}$

$c(\text{NaCl}) = \frac{0.6325}{\left(\frac{100}{1000}\right)} = 6.3 \text{ M}$

e $n(\text{NaCl}) = 6.3 \times 0.193 = 1.2 \text{ mol}$

$m(\text{NaCl}) = 1.2 \times 58.5 = 71 \text{ g}$

f Full equation: $NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

Spectator ions: $Na^+_{(aq)}$ and $NO_3^-{}_{(aq)}$

Precipitate: $AgCl_{(s)}$

Ionic equation: $Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$

g [The solution will start to turn cloudy[1]] [and a sediment of the precipitate will start to form at the bottom of the beaker.[2]]

✓ ✗ I have identified one observation of precipitate formation.[1]

✓ ✗ I have identified another observation of precipitate formation.[2]

Key science skills questions

13 a Independent variable: Temperature of solution

Dependent variable: Solubility of helium gas in solution

b Some controlled variables include:

- Pressure of helium
- Amount of helium
- Volume of water
- Surface area of water

c [According to the MSDS, helium gas may explode.[1]] [Therefore, it is appropriate to keep it in a sealed vessel and away from any external sources of ignition or heat.[2]]

✓ ✗ I have interpreted the MSDS.[1]

✓ ✗ I have identified some safety considerations.[2]

d [At higher temperatures, gas molecules have more kinetic energy.[1]] [This means that the gas molecules are able to move about more and escape the solution, decreasing their solubility at higher temperatures.[2]]

✓ ✗ I have identified the effect of temperature on kinetic energy.[1]

✓ ✗ I have explained why gas solubility decreases with temperature.[2]

8A An introduction to acids and bases

Theory review questions

1 B. According to the Brønsted-Lowry theory of acids and bases, an acid-base reaction is defined as a reaction involving the transfer of a proton, or H^+ ion.

2 A. During an acid-base reaction, a H^+ ion is transferred from a donating acid to an accepting base.

3 B. $HCl_{(aq)}$ donates a H^+ ion to form $Cl^-_{(aq)}$, therefore acting as an acid. $H_2O_{(l)}$ accepts a H^+ ion to form $H_3O^+_{(aq)}$, therefore acting as a base.

4 B. HF has one proton available to donate, making it a monoprotic acid. HF cannot also accept a H^+ ion and so is not amphiprotic as well.

5 B. $H_2PO_4^-$ has two available H^+ ions to donate and so is, therefore, a diprotic acid. It can also accept a H^+ ion to form H_3PO_4, making it amphiprotic as well.

6 A. If an acid is polyprotic, it is capable of donating more than one H^+ ion.

7 A. Species that differ by one H^+ in an acid-base reaction (formed from the acceptance or donation of one H^+ ion) such as $H_2CO_{3(aq)}$ / $HCO_3^-{}_{(aq)}$ are classified as conjugate acid-base pairs.

8 B. CH_3COOH, H_2CO_3 and NH_4^+ (ammonium) are common weak acids.

9 A. HCl, as a strong acid, will ionise (almost) completely in water, readily donates its proton and breaks up into $H^+_{(aq)}$ and $Cl^-_{(aq)}$. As a result, almost zero molecules of HCl will remain in the solution.

10 B. The given solution has only a few molecules present, meaning that it is most likely a dilute solution.Complete ionisation is characteristic of strong acids.

Deconstructed questions

11 A **12** B

13 **a** $HNC(NH_2)_{2(aq)} + H_2O_{(l)} \rightarrow C(NH_2)_3^+{}_{(aq)} + OH^-_{(aq)}$

b [The species acting as an acid is $H_2O_{(l)}$.[1]][This is because $H_2O_{(l)}$ donates a H^+ ion to form $OH^-_{(aq)}$, which is consistent with the Brønsted-Lowry theory of acids and bases that classifies an acid as a proton donor.[2]]

- ✓ ✗ I have identified the chemical species that is acting as an acid.[1]
- ✓ ✗ I have justified my answer with reference to the Brønsted-Lowry theory of acids and bases.[2]

c [$H_2O_{(l)}$ is classified as amphiprotic.[1]][As $H_2O_{(l)}$ is capable of both accepting a H^+ ion to form $H_3O^+_{(aq)}$ and donating a H^+ ion to form $OH^-_{(aq)}$, it is classified as an amphiprotic species.[2]]

- ✓ ✗ I have stated how the chemical species is classified.[1]
- ✓ ✗ I have explained my answer with reference to the relevant theory.[2]

Exam-style questions

Within lesson

14 D **15** B **16** D

17 **a** Monoprotic acid

b $HClO_{4(aq)} + NH_{3(aq)} \rightarrow ClO_4^-{}_{(aq)} + NH_4^+{}_{(aq)}$

c **i** ClO_4^-

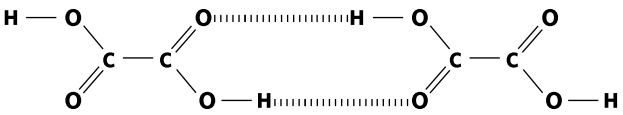
ii [$ClO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HClO_{4(aq)}$[1]][In order for the concentration of $OH^-_{(aq)}$ ions to increase, the conjugate base $ClO_4^-{}_{(aq)}$ will have to act as a base and accept a proton donated by $H_2O_{(l)}$, as shown in the given reaction.[2]][However, as the conjugate base of the very strong acid $HClO_{4(aq)}$, $ClO_4^-{}_{(aq)}$ will be a very weak base.[3]][This means that this reaction will not occur to any meaningful extent and the concentration of $OH^-_{(aq)}$ ions will stay relatively constant.[4]]

- ✓ ✗ I have included a balanced acid-base reaction equation in my answer.[1]
- ✓ ✗ I have described the acid-base reaction that would have to occur.[2]
- ✓ ✗ I have referenced the theory of the relative strengths of conjugate acid-base pairs.[3]
- ✓ ✗ I have justified my answer with reference to the included reaction and the relative strengths of conjugate acid-base pairs.[4]

18 **a** C **b** B

Multiple lessons

19 **a**



b **i** The table of values given does not include the presence of $H_2O_{(l)}$, which will be a significant proportion of the mass of the solution.

ii [Since the compounds $(COOH)_2$, $(COOH)(COO)^-$ and $(COO)_2^{2-}$ are all very similar in molar mass, their relative % composition by mass in the solution gives an indication of their concentrations.[1]][The most highly concentrated species is $(COOH)_2$ at 25.7%, with the deprotonated forms of oxalic acid only occupying 4.30% and 0.912% of the solution's mass.[2]][This means that only a small proportion of oxalic acid is deprotonated at all and as a result it can be classified as a weak acid.[3]]

- ✓ ✗ I have described how % composition by mass can indicate concentration.[1]
- ✓ ✗ I have made reference to the data present in the table.[2]
- ✓ ✗ I have used these data to justify my answer with reference to appropriate theoretical concepts.[3]

c **i** $n((COOH)_2) = 2.30 \times 5.00 = 11.5$ mol

$N((COOH)_2) = 11.5 \times 6.02 \times 10^{23} = 6.92 \times 10^{24}$ molecules

ii $n((COOH)_2) = 2.30 \times 5.00 = 11.5$ mol

$n(OH^-) = \frac{2}{1} \times n((COOH)_2) = 2 \times 11.5 = 23.0$ mol

$m(OH^-) = 23.0 \times (16.0 + 1.0) = 391$ g

iii Diprotic acid

d (1) $(COOH)(COO)^-_{(aq)}$

Conjugate acid: $(COOH)_{2(aq)}$ Conjugate base: $(COO)_2^{2-}{}_{(aq)}$

(2) $H_2O_{(l)}$

Conjugate acid: $H_3O^+_{(aq)}$ Conjugate base: $OH^-_{(aq)}$

(3) $OH^-_{(aq)}$

Conjugate acid: $H_2O_{(l)}$ Conjugate base: $O^{2-}_{(aq)}$

Key science skills

20 a i Independent variable: Identity of the acid (HF/HCl)

Dependent variable: Final concentration of H_3O^+ in the solution

ii Possible controlled variables include:

- Consistent environmental temperature
- Make-up of H_2O solution (no contaminants or foreign species)
- Time for which the pH probe is recording
- Volume of H_2O used
- Concentration of each acid (same number of mol of each species in the solution)
- Proticity of each acid

b i $n(HF) = \frac{5.00}{19.0 + 1.0} = 0.250$ mol

$n(HCl) = \frac{9.125}{35.5 + 1.0} = 0.250$ mol

$[HF] = [HCl] = \frac{0.250}{0.250} = 1.00$ M

[Since this experiment is only investigating the strength of each acid, it is important that as many other variables as possible are controlled for, including concentration.[1]][Therefore, the two masses of $HF_{(s)}$ and $HCl_{(s)}$ are different, but the number of mol is equal (0.250 mol), leading to solutions of equal concentration (1.00 M).[2]]

✓ ✗ I have included a calculation in my answer.

✓ ✗ I have stated the importance of controlling for concentration.[1]

✓ ✗ I have justified my answer with reference to my calculations.[2]

ii [Repeating an experiment reduces the effects of random fluctuations in experimental results (variance).[1]][Averaging multiple trials leads to a final result that is closer to the true value (more accurate) as long as there are no systematic errors present in the method.[2]]

✓ ✗ I have described how repetition reduces the effect of variance in results.[1]

✓ ✗ I have described the impact on the accuracy of results obtained.[2]

c [Caoimhe has collected data on the concentration of $H_3O^+_{(aq)}$ ions present in each acid's solution over time. The point at which the curve plateaus (is constant) will give the final concentration of $H_3O^+_{(aq)}$ ions in the solution.[1]][Since both HF and HCl are monoprotic acids, this final concentration of $H_3O^+_{(aq)}$ can be used to determine what percentage of HCl or HF ionised into $H^+_{(aq)}$ ($H_3O^+_{(aq)}$) ions, which is an indicator of each acid's strength.[2]] [The final concentrations of $H_3O^+_{(aq)}$ were 0.025 M (HF) and 1.0 M (HCl), making hydrochloric acid significantly stronger than hydrofluoric acid given that they both started at the same concentration (1.00 M).[3]]

✓ ✗ I have described how to analyse the data collected.[1]

✓ ✗ I have explained how the data collected can be used to determine each acid's strength.[2]

✓ ✗ I have identified which acid is stronger with justification from the data collected.[3]

d Acid-resistant personal protective equipment (gloves, protective eyewear, aprons) should be worn and the HF beaker should be made of a HF-resistant material other than glass.

e [Unlike HF and HCl, phosphoric acid is a triprotic acid and so will deprotonate in three stages, each of which will give off a proton to form $H_3O^+_{(aq)}$ ions.[1]][However, the current experimental technique only measures the total hydronium concentration, not from where those ions have come.[2]][This means that a H_3PO_4 molecule that has donated two or three protons will contribute twice or three times as much respectively as one that has only lost one proton. This will skew the data and give an incorrect estimation of the acid's strength.[3]] [Since there is an extra variable here (the proticity of the acid) affecting the results (dependent variable), the experiment is no longer considered valid.[4]]

✓ ✗ I have identified the key difference between H_3PO_4, HF and HCl.[1]

✓ ✗ I have identified the issue with the experimental technique.[2]

✓ ✗ I have explained why this key difference will result in inaccurate data being collected.[3]

✓ ✗ I have explained how this will affect the validity of the experiment.[4]

8B Measuring acidity

Theory review questions

1 B. The expression $K_W = [H_3O^+][OH^-] = 10^{-14}$ M^2 is constant at 25°C. In a solution of pure water, $[H_3O^+] = [OH^-] = 10^{-7}$ M.

2 A. For a constant temperature, the ionic product of water is constant, but the concentrations of H_3O^+ and OH^- will adjust whenever an acid or base is added to the solution.

3 A. In an acidic solution the concentration of H_3O^+ ions is greater than the concentration of OH^- ions. In a basic solution, the opposite is true.

4 B. A pH value of 7 is considered neutral. pH values above or below neutral are referred to as being more basic and acidic respectively.

5 A. Indicators such as methyl orange are compounds that undergo colour changes as the pH of a solution changes.

6 B. Whole number pH values differ from each other by a factor of 10.

7 A. Increasing the concentration of an acid in solution will also increase the concentration of H_3O^+ ions.

8 B. Adding pure water (not an acid or base) to a solution is also known as diluting the solution.

9 B. Diluting an acidic solution (adding pure water) will decrease the concentration of H_3O^+ ions in the solution.

10 A. Ammonia is a weak base. Diluting a basic solution (adding pure water) will decrease the concentration of OH^- ions in the solution and increase the concentration of H_3O^+ ions.

Deconstructed questions

11 C **12** B

13 Initial concentration of $HNO_{3(aq)}$ = 0.78 M

Diluted solution volume: 9.3 + 25.0 = 34.3 L

$c_1V_1 = c_2V_2$

$0.78 \times 9.3 = c_2 \times 34.3$

$c_2 = \frac{0.78 \times 9.3}{34.3} = 0.2115$ M

$[H_3O^+] = [HNO_3] = 0.2115$ M

$pH = -\log_{10}[H_3O^+] = -\log_{10}(0.2115) = 0.67$

Exam-style questions

Within lesson

14 B **15** D

16 **a** $HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

$[HCl] = [H_3O^+] = 0.31$ M

$pH = -\log_{10}(0.31) = 0.51$

b $H_2SO_{4(aq)} + 2H_2O_{(l)} \rightarrow 2H_3O^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

Diluted volume = 30.0 + 12.0 = 42.0 L

$c_1V_1 = c_2V_2$

$2.08 \times 30.0 = c_2 \times 42.0$

$[H_2SO_4]_{diluted} = c_2 = 1.4857$ M

$[H_3O^+] = 2 \times [H_2SO_4]$

$[H_3O^+] = 2 \times 1.4857 = 2.971$ M

$pH = -\log_{10}(2.971) = -0.473$

c $HNO_{3(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$

0.11 kg L^{-1} = 110 g L^{-1}

$n(HNO_3) = \frac{110}{1.0 + 14.0 + (3 \times 16.0)} = 1.746$ mol

$c(HNO_3) = \frac{1.746}{1} = 1.746$ M

Diluted volume = 640 + 90.4 = 730.4 L

$1.746 \times 640 = c_2 \times 730.4$

$[HNO_3]_{diluted} = c_2 = 1.53$ M

$[H_3O^+] = [HNO_3] = 1.53$ M

$pH = -\log_{10}(1.53) = -0.18$

d [The ionic product of water, K_W, is a constant value equal to $[H_3O^+] \times [OH^-] = 10^{-14}$ M^2.[1]][As the concentration of H_3O^+ ions increases, the concentration of OH^- ions will decrease to maintain its constant value.[2]][The lower a solution's pH, the higher its concentration of H_3O^+ ions. Therefore, the solution of part **b** will have the lowest concentration of OH^- ions since it has the lowest pH.[3]]

✓ ✗	I have defined the ionic product of water.[1]
✓ ✗	I have described how K_W adjusts depending on relative ion concentrations.[2]
✓ ✗	I have linked my answer to the question and pH values.[3]

e **i** Basic

ii Highest: pH = 7.35

$[H_3O^+] = 10^{-7.35} = 4.47 \times 10^{-8}$ M

Lowest: pH = 7.45

$[H_3O^+] = 10^{-7.45} = 3.55 \times 10^{-8}$ M

17 **a** $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

$[NaOH] = [OH^-] = 0.4$ M

$[H_3O^+] \times 0.4 = 10^{-14}$

$[H_3O^+] = \frac{10^{-14}}{0.4} = 2.5 \times 10^{-14}$ M

$pH = -\log_{10}(2.5 \times 10^{-14}) = 13.6$

b **i** $Ca(OH)_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2OH^-_{(aq)}$

$n(Ca(OH)_2) = \frac{8.3}{40.1 + 2 \times (16.0 + 1.0)} = 0.112$ mol

$[Ca(OH)_2] = \frac{0.112}{0.324} = 0.3457$ M

$[OH^-] = 2 \times [Ca(OH)_2] = 2 \times 0.3457 = 0.6914$ M

$[H_3O^+] \times 0.6914 = 10^{-14}$

$[H_3O^+] = \frac{10^{-14}}{0.6914} = 1.4464 \times 10^{-14}$ M

$pH = -\log_{10}(1.4464 \times 10^{-14}) = 13.84$

ii For a pH of 9.20: $[H_3O^+] = 10^{-pH} = 10^{-9.20} = 6.31 \times 10^{-10}$ M

$[OH^-]_{diluted} = \frac{10^{-14}}{6.31 \times 10^{-10}} = 1.585 \times 10^{-6}$ M

$[OH^-]_{initial} = 0.6914$ M

$0.6914 \times 0.324 = 1.585 \times 10^{-6} \times V_2$

$V_2 = 14134$ L

Volume added = 14134 − 0.324 = 1.4×10^4 L

c $KOH_{(s)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$

$n(KOH) = cV + n(KOH)_{added} = 0.90 \times 27.0 + 4.0 = 28.3$ mol

$[KOH] = \frac{28.3}{27.0} = 1.048$ M

$[OH^-] = [KOH] = 1.048$ M

$1.048 \times 27.0 = c_2 \times (27.0 + 40.4)$

$[OH^-]_{diluted} = c_2 = 0.4198$ M

$[H_3O^+] \times 0.4198 = 10^{-14}$

$[H_3O^+] = \frac{10^{-14}}{0.4198} = 2.382 \times 10^{-14}$ M

$pH = -\log_{10}(2.382 \times 10^{-14}) = 13.62$

d **i** $[H_3O^+] \times 0.4198 = 51.3 \times 10^{-14}$

$[H_3O^+] = \frac{51.3 \times 10^{-14}}{0.4198} = 1.222 \times 10^{-12}$ M

$pH = -\log_{10}(1.222 \times 10^{-12}) = 11.91$

The pH of the solution would be lower (more acidic).

ii At neutral pH, $[H_3O^+] = [OH^-] = a$

$K_W = a \times a = 51.3 \times 10^{-14}$

$a = \sqrt{51.3 \times 10^{-14}} = 7.162 \times 10^{-7} = [H_3O^+]$

$pH = -\log_{10}(7.162 \times 10^{-7}) = 6.145$

Multiple lessons

18 **a** Diprotic acid

b **i** $n(H_2SeO_3) = 3.9$ mol

$n(HSeO_3^-) = 3.9 \times 0.65 = 2.535$ mol

$n(SeO_3^{2-}) = 3.9 \times 0.15 = 0.585$ mol

$n(H^+) = n(HSeO_3^-) + 2 \times n(SeO_3^{2-}) = 2.535 + 2 \times 0.585$

$n(H^+) = 3.7$ mol

ii $n(H_3O^+) = n(H^+) = 3.705$ mol

$V = 72$ L

$[H_3O^+] = \frac{3.705}{72} = 0.05146$ M

$pH = -\log_{10}(0.05146) = 1.29$

c [The deprotonated selenous acid, $HSeO_3^-$, has a negative charge that can be thought of as electrostatically attracting and 'holding on' to the positive H^+ ion.[1]][This means that it becomes more difficult for the positive H^+ ion to escape the pull of the remaining compound's negative charge, resulting in fewer H_2SeO_3 molecules that have successfully lost both protons.[2]]

✓ ✗ I have explained the relevant theory about deprotonated polyprotic acids.[1]

✓ ✗ I have linked my explanation to the question's scenario.[2]

d [If selenous acid were a stronger acid, it would donate its protons more readily.[1]][This would lead to a greater amount of H_3O^+ ions in the solution as more of the H_2SeO_3 deprotonates either once or twice.[2]][A higher concentration of H_3O^+ ions in the solution would correspond to a lower pH.[3]]

✓ ✗ I have described how the strength of an acid and its willingness to donate protons are related.[1]

✓ ✗ I have explained how this will affect the $[H_3O^+]$ in the solution.[2]

✓ ✗ I have explained how this will affect the pH of the solution.[3]

19 **a** Blue

b [According to the Brønsted-Lowry theory of acids and bases, an acid donates a proton to an accepting base.[1]][In this reaction bromophenol blue, $C_{19}H_{10}Br_4O_5S_{(aq)}$, has donated a proton to form $C_{19}H_9Br_4O_5S^-_{(aq)}$, meaning that it acted as an acid in this reaction.[2]]

✓ ✗ I have stated the definition of an acid or base.[1]

✓ ✗ I have explained whether bromophenol blue would be an acid or base in the given reaction.[2]

c For pH 2.00: $[H_3O^+]_1 = 10^{-2.00}$, $V_1 = 0.0300$ L

For pH 4.60 (blue): $[H_3O^+]_2 = 10^{-4.60}$

$10^{-2.00} \times 0.0300 = 10^{-4.60} \times V_2$

$V_2 = 11.94$ L

Volume added = 11.94 − 0.0300 = 12 L

d pH = 4.60, $[H_3O^+] = 10^{-4.60}$

$10^{-4.60} \times [OH^-] = 10^{-14}$

$[OH^-] = \frac{10^{-14}}{10^{-4.60}} = 4.0 \times 10^{-10}$ M

Key science skills

20 **a** The beaker of deionised water acts as a control for this experiment.

b Independent variable: Identity of monoprotic acid

Dependent variable: Brightness of the lamp (in lumens)

c [Solutions 1–3 are acidic with a pH below 7, whereas solution 4 is water, which will have the highest pH of 7.[1]][Electrical conductivity increases with ion concentration and so solutions with a high brightness will have a high concentration of ions (H_3O^+ and the acid's conjugate base).[2]][A high concentration of H_3O^+ also implies a low pH and so an increase in brightness will correspond to a decrease in pH.[3]][Since high brightness corresponds to low pH, in order of decreasing pH: 4, 1, 3, 2.[4]]

✓ ✗ I have stated the pH of beaker 4 and how the other beakers relate to it.[1]

✓ ✗ I have described the relationship between electrical conductivity, ion concentration and brightness.[2]

✓ ✗ I have explained the relationship between brightness and pH.[3]

✓ ✗ I have used this theory to list the solutions in beakers 1–4 in order of decreasing pH.[4]

d **i** If the concentrations of each of the solutions in beakers 1–3 are all increased, then the brightness of each lamp will also increase.

ii Quantitative

8C Reactions of acids

Theory review questions

1 B. Spectator ions are chemical species that are present as ions on both sides of a chemical equation and so they are not included in the ionic equation.

2 B. Water molecules are covalent compounds and so must be included in ionic equations.

3 A. Acids and reactive metals react together to form an ionic salt, as well as hydrogen gas.

4 A. Ionic salts contain a positively charged cation (Na^+) in an ionic bond with a negatively charged anion (Cl^-). CH_4 is a covalent substance and HCl is an acid.

5 B. The general equation for this reaction is:
Acid + Metal hydroxide → Ionic salt + $H_2O_{(l)}$.

6 B. Spectator ions are chemical species that are present as ions at the start and end of a chemical reaction. They will not form covalent bonds and will remain in an aqueous state.

7 B. When acids react with metal hydroxides, the $H^+_{(aq)}$ and $OH^-_{(aq)}$ ions react to form neutral water, $H_2O_{(l)}$.

8 A. Metal carbonates are compounds of a metal cation and the carbonate ion, CO_3^{2-}. When reacting with acids, they will produce an ionic salt, carbon dioxide gas and liquid water.

9 B. In the reaction between acids and soluble metal hydroxides, only the H^+ and OH^- ions are reacting together.

10 B. When acids react with metal carbonates, an ionic salt, carbon dioxide gas and liquid water will be produced.

Deconstructed questions

11 B **12** D

13 **a** $I^-_{(aq)}$ and $K^+_{(aq)}$

b $2H^+_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

Exam-style questions

Within lesson

14 D **15** C

16 **a** $CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow CH_3COONa_{(aq)} + H_2O_{(l)}$

b $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

c [$CH_3COO^-_{(aq)}$ and $Na^+_{(aq)}$ are present as ions throughout the whole reaction and so are not included in the ionic equation as they are spectator ions.[1]][The $H^+_{(aq)}$ ion from ethanoic acid and the $OH^-_{(aq)}$ ion from sodium hydroxide combine to form a covalent water molecule, $H_2O_{(l)}$, and so are included in the ionic equation.[2]]

✓ ✗ I have explained my reasoning for the omission of ions in the ionic equation.[1]

✓ ✗ I have explained my reasoning for the inclusion of ions in the ionic equation.[2]

d Neutralisation reaction

Multiple lessons

17 **a** $H_2SO_{4(aq)} + ZnCO_{3(s)} \rightarrow ZnSO_{4(aq)} + CO_{2(g)} + H_2O_{(l)}$

b $2H^+_{(aq)} + ZnCO_{3(s)} \rightarrow Zn^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

c [In this reaction, carbon dioxide gas is produced from the zinc carbonate.[1]][This gas will leave the solution and spread into the atmosphere, decreasing the volume (water level) of the solution in the beaker.[2]]

✓ ✗ I have identified the gas that is evolved in this reaction.[1]

✓ ✗ I have explained how this results in the observed change to the water level.[2]

d The difference in masses will be the mass of $CO_{2(g)}$ that is produced.

$H_2SO_{4(aq)} + ZnCO_{3(s)} \rightarrow ZnSO_{4(aq)} + CO_{2(g)} + H_2O_{(l)}$

Once all of the $ZnCO_{3(s)}$ has reacted, no more $CO_{2(g)}$ can be produced.

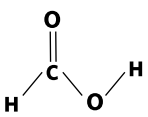
$n(ZnCO_{3(s)}) = \dfrac{40.3}{65.4 + 12.0 + 3 \times 16.0} = 0.3214$ mol

$n(CO_{2(g)}) = n(ZnCO_{3(s)}) = 0.3214$ mol

$m(CO_2) = (12.0 + 2 \times 16.0) \times 0.3214 = 14.14$ g

Therefore, the change in mass would be a decrease of 14 g.

18 **a**



b carboxyl

c **i** $2HCOOH_{(aq)} + Sn_{(s)} \rightarrow (HCOO)_2Sn_{(aq)} + H_{2(g)}$

ii $2H^+_{(aq)} + Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + H_{2(g)}$

d $2HCOOH_{(aq)} + Sn_{(s)} \rightarrow (HCOO)_2Sn_{(aq)} + H_{2(g)}$

$n(Sn_{(s)}) = \dfrac{40.0}{118.7} = 0.337$ mol

$n(HCOOH_{(aq)}) = \dfrac{2}{1} \times n(Sn_{(s)}) = 0.674$ mol

$m(HCOOH_{(aq)}) = 0.674 \times 46.0 = 31.0$ g

Number of ants $= \dfrac{31.0}{0.600 \times 10^{-6}} = 5.17 \times 10^7$ ants

e $n(Sn_{(s)}) = 0.337$ mol

$n(H_{2(g)}) = \dfrac{1}{1} \times n(Sn_{(s)}) = \dfrac{1}{1} \times 0.337 = 0.337$ mol

$N(H_{2(g)}) = 0.337 \times 6.02 \times 10^{23} = 2.03 \times 10^{23}$ molecules of $H_{2(g)}$

19 **a** $HCl_{(aq)} + H_2O_{(l)} \rightarrow Cl^-_{(aq)} + H_3O^+_{(aq)}$

$[HCl] = [H_3O^+]$ (1 : 1 ratio)

$pH = -\log_{10}(3.2 \times 10^{-4}) = 3.49$

b monoprotic

c **i** $3HCl_{(aq)} + Al(OH)_{3(s)} \rightarrow AlCl_{3(aq)} + 3H_2O_{(l)}$

ii $3H^+_{(aq)} + Al(OH)_{3(s)} \rightarrow Al^{3+}_{(aq)} + 3H_2O_{(l)}$

d $3HCl_{(aq)} + Al(OH)_{3(s)} \rightarrow AlCl_{3(aq)} + 3H_2O_{(l)}$

$[H^+]_{healthy} = 3.2 \times 10^{-4}$ M

$n(H^+)_{healthy} = 3.2 \times 10^{-4} \times 1.20 = 3.84 \times 10^{-4}$ mol

$[H^+]_{acid\ reflux} = 10^{-1.20}$ M

$n(H^+)_{acid\ reflux} = 10^{-1.20} \times 1.20 = 7.57 \times 10^{-2}$ mol

$n(H^+)_{acid\ reflux} - n(H^+)_{healthy} = (7.57 \times 10^{-2}) - (3.84 \times 10^{-4})$

$= 7.533 \times 10^{-2}$ mol

Therefore, 7.533×10^{-2} mol of H^+ must be consumed.

$n(Al(OH)_{3(s)})_{required} = \dfrac{1}{3} \times 7.533 \times 10^{-2} = 2.51 \times 10^{-2}$ mol

$m(Al(OH)_{3(s)})_{required} = 2.51 \times 10^{-2} \times 78 = 1.96$ g

$N(Al(OH)_{3(s)}$ tablets$) = \dfrac{1.96}{250 \times 10^{-3}} = 7.83$ tablets

8 tablets would be required to return the stomach to average pH.

Key science skills

20 **a** [When $MgCO_{3(s)}$ reacts with the nitric acid, $CO_{2(g)}$ will be produced.[1]][This $CO_{2(g)}$ will rise out of the test tube with $MgCO_{3(s)}$ and deposit in the test tube with $Ca(OH)_{2(aq)}$ (limewater), causing the limewater to become a milky/cloudy colour.[2]][The only test tube that has caused this reaction with the limewater is test tube 1, which must, therefore, be the test tube containing $MgCO_{3(s)}$.[3]]

✓ ✗ I have identified the key product from the reaction between $MgCO_{3(s)}$ and nitric acid.[1]

✓ ✗ I have explained how this product will affect the results of the investigation.[2]

✓ ✗ I have used this information to answer the question.[3]

b [When $Pb_{(s)}$ reacts with nitric acid, $H_{2(g)}$ is produced, but not $CO_{2(g)}$.[1]][This can be observed as the bubbling/fizzing present in test tube 5, but the absence of a reaction with the limewater/$Ca(OH)_{2(aq)}$ in test tube 6.[2]][When $Zn(OH)_{2(s)}$ reacts with nitric acid, no noticeable products would be observed as the ionic salt $Zn(NO_3)_2$ is soluble in water.][This is consistent with the observed results for test tube 3.[3]][Therefore, test tube 5 contains $Pb_{(s)}$ and test tube 3 contains $Zn(OH)_{2(s)}$.[4]]

✓ ✗ I have identified the key product of the reaction between nitric acid and $Pb_{(s)}$.[1]

✓ ✗ I have explained how this would impact the observed results for $Pb_{(s)}$.[2]

✓ ✗ I have identified the key product of the reaction between nitric acid and $Zn(OH)_{2(s)}$.[3]

✓ ✗ I have used linked my explanation to the question.[4]

c [The data collected are related to the presence/absence of an observed change (fizzing/no fizzing and milky limewater/clear solution) not numerical descriptions of these data (eg. number of bubbles evolved per second or light absorbance of the limewater test tube).[1]] [Therefore, the data collected are qualitative.[2]]

✓ ✗ I have explained the nature of the data collected.[1]

✓ ✗ I have classified the data as qualitative or quantitative.[2]

d 1. Personal Protective Equipment (PPE) such as gloves or glasses to guard against potential minor explosions

2. Keep any open sources of flame or heat far away from the hydrogen gas.

8D Acid-base chemistry in the body

Theory review questions

1 A. Amino acids contain two functional groups: an amino group ($-NH_2$) and a carboxyl group (-COOH).

2 B. The amino group can accept a proton and the carboxyl group can donate a proton.

3 B. Individual amino acids can be linked together by peptide/amide bonds with molecular formula CONH.

4 B. The amino group has accepted a proton to form $-NH_3^+$, meaning the amino acid is acting as a Brønsted-Lowry base.

5 A. Enzymes are proteins (chains of amino acids) that act to increase the rate of chemical reactions.

6 B. The rate at which an enzyme converts reactants to products (its activity) depends heavily on pH due to their acid-base properties.

7 B. Carbonic anhydrase increases the rate of reaction between gaseous carbon dioxide and water and unionised carbonic acid.

8 B. Most CO_2 is converted to H_2CO_3 by carbonic anhydrase, which then ionises to allow it to be transported as highly soluble bicarbonate ions, HCO_3^-.

9 A. Hypoventilation is a decrease in breathing rate, whereas hyperventilation is an increase.

10 B. Increased breathing out of $CO_{2(g)}$ during hyperventilation reduces the amount of $CO_{2(g)}$ in the bloodstream available to be converted to carbonic acid. This leads to an increased blood pH, or alkalosis.

Deconstructed questions

11 D **12** C

13 [The decrease in breathing rate due to COPD would result in a higher amount of $CO_{2(g)}$ dissolved in the blood.[1]] [This increase in $CO_{2(g)}$ is then converted to carbonic acid by carbonic anhydrase according to the equation $CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$.[2]] [This leads to a higher amount of carbonic acid and, therefore, a higher concentration of $H^+_{(aq)}$ in the blood as carbonic acid ionises according to the equation $H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)}$.[3]] [The increased $[H^+]$ results in a decreased blood pH.[4]]

✓ ✗ I have described the effect of COPD on blood $CO_{2(g)}$.[1]

✓ ✗ I have explained what happens to $CO_{2(g)}$ in the blood, with reference to a relevant equation.[2]

✓ ✗ I have explained what happens to $H_2CO_{3(aq)}$ in the blood, with reference to a relevant equation.[3]

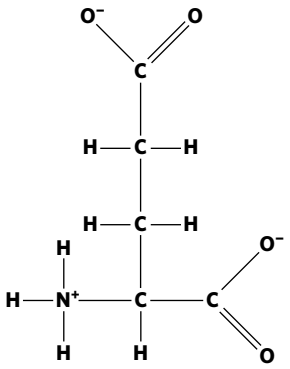
✓ ✗ I have linked my answer to the effect on blood pH.[4]

Exam-style questions

Within lesson

14 D **15** B

16 **a**



b [This molecule is best classified as amphiprotic.[1]] [This is because the $-NH_3^+$ group can donate a proton and the two $-COO^-$ groups are capable of accepting protons.[2]]

✓ ✗ I have classified the molecule as polyprotic, monoprotic or amphiprotic.[1]

✓ ✗ I have explained my answer with reference to the functional groups present in the molecule.[2]

c −1

17 **a** Hyperventilation

b [$CO_{2(g)}$ is breathed out at a higher rate during hyperventilation.[1]] [This means that a lower amount of $CO_{2(g)}$ remains in the body, leading to a decrease in the amount of $CO_{2(g)}$ dissolved in the bloodstream.[2]]

✓ ✗ I have described how hyperventilation influences how quickly $CO_{2(g)}$ leaves the body.[1]

✓ ✗ I have explained how hyperventilation will lead to a decrease in blood $CO_{2(g)}$.[2]

c [The decrease in blood $CO_{2(g)}$ lowers the amount of $CO_{2(g)}$ available to be converted into carbonic acid by carbonic andhydrase.[1]] [This leads to a decrease in the amount of carbonic acid that can ionise to form $H^+_{(aq)}$ ions, resulting in an increased blood pH.[2]]

✓ ✗ I have described the role of carbonic anhydrase in regulating blood pH.[1]

✓ ✗ I have used this information to explain how this change in blood $CO_{2(g)}$ will affect blood pH.[2]

d Alkalosis

Multiple lessons

18 a The weak acid $H_2CO_{3(aq)}$ will have a strong conjugate base, $HCO_3^-{}_{(aq)}$.

b [Since the bicarbonate ion, $HCO_3^-{}_{(aq)}$ is a relatively strong base, carbonic acid will not deprotonate further to any great extent as bases accept protons, rather than donate.[1]] [Therefore, this second deprotonation will have essentially no impact on the pH of the blood.[2]]

✓ ✗ I have described the extent of the bicarbonate ion's deprotonation.[1]

✓ ✗ I have explained how this will impact the pH of the blood.[2]

19 a Hydrogen bonding, permanent dipole-dipole forces, dispersion forces

b [Both serine and ethanol are highly polar molecules due to the presence of a number of highly electronegative atoms and functional groups.[1]] [Therefore, the polar molecule serine will dissolve in the beaker of the similarly polar molecule ethanol.[2]]

✓ ✗ I have described the polarity of serine and ethanol.[1]

✓ ✗ I have explained how their polarity will affect their solubility.[2]

c i Amphiprotic

ii $C_3H_7NO_{3(aq)} + H_2O_{(l)} \rightarrow C_3H_6NO_3^-{}_{(aq)} + H_3O^+_{(aq)}$

iii Conjugate acid/conjugate base:

$C_3H_7NO_{3(aq)}/C_3H_6NO_3^-{}_{(aq)}$

$H_3O^+_{(aq)}/H_2O_{(l)}$

Key science skills

20 a Independent variable: pH of solution (due to concentration of $NaOH_{(aq)}/HCl_{(aq)}/H_2O_{(l)}$)

Dependent variable: mL of gas in syringe ($CO_{2(g)}$)

b [Gaseous carbon dioxide bubbled through the solution in the conical flask can be converted to aqueous carbonic acid by the enzyme carbonic anhydrase.[1]] [Any $CO_{2(g)}$ that is not converted to $H_2CO_{3(aq)}$ and fixed in the solution, will just bubble through, rise and enter the syringe.[2]] [Therefore, the level to which the syringe expands will increase with the amount of $CO_{2(g)}$ that escapes the solution.[3]] [The amount of $CO_{2(g)}$ that is converted to $H_2CO_{3(aq)}$ will depend on the activity of the enzyme, meaning that a decrease in the activity of the enzyme will result in an increase in the expansion of the syringe.[4]]

✓ ✗ I have described the reaction that would take place within the conical flask.[1]

✓ ✗ I have explained the link between this reaction and expansion of the syringe.[2]

✓ ✗ I have explained the link between expansion of the syringe and the amount of $CO_{2(g)}$ that escapes.[3]

✓ ✗ I have answered the question by explaining the link between enzyme activity and the volume of the syringe.[4]

c [The activity of carbonic anhydrase is highest at a pH of approximately 7.5.[1]] [This is consistent with what would be expected of an enzyme present in the human bloodstream, as human blood pH is tightly regulated around a pH of 7.35–7.45.[2]]

✓ ✗ I have identified the pH at which the enzyme's activity is highest.[1]

✓ ✗ I have explained whether this aligns with the pH of the human bloodstream.[2]

Chapter 8 review

Multiple choice questions

1 B. Brønsted-Lowry acid-base reactions involve the transfer of a proton (H^+) from an acid to a base.

2 C. $HPO_4^{2-}{}_{(aq)}$ is an amphiprotic species.

3 C. Ionic equations do not include spectator ions and will only include chemical species that undergo a chemical change.

4 A. $H_2PO_4^-{}_{(aq)}$ is an amphiprotic species as it can both donate and accept protons and is polyprotic as it can donate two protons.

5 B. Pure water has a pH of 7. Ethanoic acid is a weak acid and so the pH will decrease from 7 to a value significantly higher than 1.23.

6 B. $H_2SeO_{3(aq)}$ is a Brønsted-Lowry acid that donates a proton to form its conjugate base, $HSeO_3^-{}_{(aq)}$.

7 D. Turkey and human vulture stomachs may just have different concentrations of the same acid (1000 times more concentrated), not necessarily different acids. A lower pH corresponds to a higher concentration of $H^+_{(aq)}$ ions.

8 D. Hyperventilation causes less $CO_{2(g)}$ to remain in the body and so less $CO_{2(g)}$ is converted to carbonic acid by carbonic anhydrase, increasing blood pH.

9 B. The reactive metal magnesium will react with iodic acid to form an ionic salt and hydrogen gas.

10 A. I: pH = 1.72, II: pH = 12.53, III: pH = 12.41, IV: pH = 4.43

Short answer questions

11 a $2HBr_{(aq)} + MgCO_{3(s)} \rightarrow MgBr_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

b $2H^+_{(aq)} + MgCO_{3(s)} \rightarrow Mg^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

c [The conjugate base of hydrobromic acid is $Br^-_{(aq)}$.[1]] [The stronger an acid, the weaker its conjugate base.[2]] [As $HBr_{(aq)}$ is a very strong acid, its conjugate base $Br^-_{(aq)}$ will be a base of negligible strength and so will not accept protons from water to form $OH^-_{(aq)}$ ions.[3]] [This means that the concentration of $OH^-_{(aq)}$ ions would not increase by any noticeable extent.[4]]

- ✓ ✗ I have identified the conjugate base of hydrobromic acid.[1]
- ✓ ✗ I have explained the relationship between acids and their conjugate bases.[2]
- ✓ ✗ I have explained how this relationship would affect how bromide acts in solution.[3]
- ✓ ✗ I have linked my answer to the question.[4]

d $2HBr_{(aq)} + MgCO_{3(s)} \rightarrow MgBr_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

$m(CO_{2(g)}) = 1282.3 - 1240.7 = 41.6$ g

$n(CO_{2(g)}) = \frac{41.6}{12.0 + 2 \times 16.0} = 0.945$ mol

$n(MgCO_{3(s)}) = n(CO_{2(g)}) = 0.945$ mol

$m(MgCO_{3(s)}) = 0.945 \times (24.3 + 12.0 + 3 \times 16.0) = 79.7$ g

e $[H_3O^+]_{diluted} = 10^{-5.00}$ M

$0.25 \times V_1 = 10^{-5.00} \times V_2$

$\frac{V_2}{V_1} = \frac{0.25}{10^{-5.00}} = 25\,000$

The solution would have to be diluted by a factor of 25 000.

12 a [Evan's first mistake is that the concentration of $OH^-_{(aq)}$ ions should actually be double the concentration of $Ba(OH)_{2(aq)}$ at $0.382 \times 2 = 0.76$ M.[1]][Evan has also mistakenly calculated the concentration of $H_3O^+_{(aq)}$ before the dilution calculation, as this should occur after the diluted concentration of $OH^-_{(aq)}$ is calculated.[2]] [Evan has incorrectly calculated the value of V_2, which should be $12.0 + 0.500 = 12.5$ L.[3]][Finally, Evan has used the concentration of $H_3O^+_{(aq)}$ in his dilution calculation instead of correctly using the concentration of $OH^-_{(aq)}$.[4]]

- ✓ ✗ I have identified and explained the first mistake in Evan's calculations.[1]
- ✓ ✗ I have identified and explained the second mistake in Evan's calculations.[2]
- ✓ ✗ I have identified and explained the third mistake in Evan's calculations.[3]
- ✓ ✗ I have identified and explained the fourth mistake in Evan's calculations.[4]

b $Ba(OH)_{2(aq)} \rightarrow Ba^{2+}_{(aq)} + 2OH^-_{(aq)}$

$[OH^-_{(aq)}] = 2 \times [Ba(OH)_{2(aq)}] = 2 \times 0.38 = 0.76$ M

$0.76 \times 0.500 = (12.0 + 0.500) \times c_2$

$c_2 = \frac{0.76 \times 0.500}{12.5} = 0.0304$ M

$[H_3O^+] = \frac{10^{-14}}{0.0304} = 3.29 \times 10^{-13}$ M

$pH = -\log_{10}(3.29 \times 10^{-13}) = 12.48$

c $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

d [The pH of the barium hydroxide solution will decrease[1]][as $OH^-_{(aq)}$ ions are being consumed, increasing the relative concentration of $H^+_{(aq)}$ ions.[2]]

- ✓ ✗ I have identified the change in pH of the solution.[1]
- ✓ ✗ I have identified the change in concentration of OH^- and H^+ ions.[2]

e Neutralisation reaction

13 a $CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$

b [Carbonic acid will ionise to form bicarbonate and free protons according to the equation $H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)}$.[1]] [This results in the production of $H^+_{(aq)}$ ions, which will decrease the pH of the blood.[2]][Therefore, an increase or decrease in the amount of $CO_{2(g)}$ in the blood will affect the concentration of $H^+_{(aq)}$ ions in the blood, modulating its pH.[3]]

- ✓ ✗ I have included a second chemical equation in my answer.[1]
- ✓ ✗ I have explained how this results in changes to blood pH.[2]
- ✓ ✗ I have linked my answer to the reaction from part **a**.[3]

c [The decreased breathing rate during hypoventilation results in a lower amount of $CO_{2(g)}$ leaving the body.[1]][This leads to an increase in the amount of $CO_{2(g)}$ present in the bloodstream, which is converted to carbonic acid by carbonic anhydrase.[2]][This carbonic acid then ionises to form $H^+_{(aq)}$ ions, decreasing the pH of the blood.[3]]

- ✓ ✗ I have described how hypoventilation affects carbon dioxide balance.[1]
- ✓ ✗ I have explained how this affects the amount of carbonic acid present in the blood.[2]
- ✓ ✗ I have explained how this will affect the pH of the blood.[3]

14 a $2HNO_{3(aq)} + Zn_{(s)} \rightarrow Zn(NO_3)_{2(aq)} + H_{2(g)}$

b $2H^+_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$

c [Ingot A is gold as it does not react with the nitric acid.[1]][Ingot B is Prince's metal as the zinc present in the alloy is reacting with the nitric acid to produce hydrogen gas. This is observed as the fizzing/bubbling, as well as the corrosion, of ingot B.[2]]

- ✓ ✗ I have identified the metal of ingot A and linked my answer to the experimental observations.[1]
- ✓ ✗ I have identified the metal of ingot B and linked my answer to the experimental observations.[2]

d [The pH of the solution will increase[1]][as $H^+_{(aq)}$ ions are being consumed in the reaction to form $H_{2(g)}$.[2]]

- ✓ ✗ I have identified the change in pH of the solution.[1]
- ✓ ✗ I have identified the change in H^+ ions.[2]

Key science skills questions

15 a

1. Add 500 mL of 1.0 M $NH_{3(aq)}$ to a clean, dry 750 mL glass beaker.
2. Place the pH probe in the solution so that it is fully submerged.
3. Record the pH of the solution once the reading stabilises.
4. Remove the pH probe from the solution, rinse it with deionised water and then dry it.
5. Repeat steps 1–4 with 500 mL of 1.0 M $HONH_{2(aq)}$.
6. Measure 250 mL of deionised water with the 500 mL measuring cylinder.
7. Add 250 mL of deionised water to the 250 mL of 2.0 M $N_2H_{4(aq)}$ in a clean, dry 750 mL beaker.
8. Repeat steps 2–4 with this solution.

b [The litmus paper will turn blue in all of the basic solutions and so cannot be used to determine their relative strengths.[1]][The phenolphthalein indicator can only undergo one colour change (between colourless and pink). This means that at least two of the solutions will be the same colour and their relative strength will not be able to be determined.[2]][The pH probe and recording device allow for a precise quantitative measurement of degree of ionisation of each base in solution and so can be used to determine each base's relative strength.[3]]

- ✓ ✗ I have explained why litmus paper should not be used.[1]
- ✓ ✗ I have explained why phenolphthalein indicator should not be used.[2]
- ✓ ✗ I have explained why the pH probe and recording device should be used.[3]

c [The strongest base will ionise to the greatest extent, resulting in the highest concentration of $OH^-_{(aq)}$ ions and the highest pH as measured by the pH probe.[1]][Weaker bases will ionise to a lower extent, leading to a lower concentration of $OH^-_{(aq)}$ ions and a lower pH.[2]] [This information can be used to determine which bases are stronger/weaker depending on the pH of the solutions they form.[3]]

- ✓ ✗ I have explained the link between a stronger base and pH.[1]
- ✓ ✗ I have explained the link between a weaker base and pH.[2]
- ✓ ✗ I have linked my answer to the question.[3]

d

	Ammonia	**Hydroxylamine**	**Hydrazine**
$[OH^-]$	Highest	Lowest	Middle
$[H_3O^+]$	Lowest	Highest	Middle

9A An introduction to redox reactions

Theory review questions

1 B. Redox reactions require both a reduction and oxidation half-equation to occur simultaneously.

2 B. Oxidation is the loss of electrons.

3 A. Remember the acronym OILRIG.

4 A. Redox reactions can be depicted by the movement of electrons.

5 B. In a chemical reaction, one species undergoes oxidation whilst another undergoes reduction.

6 B. Oxidising agents cause another chemical species to undergo oxidation.

7 B. Conjugate redox pairs consist of an oxidising agent and its conjugate reducing agent.

8 B. Oxidising agents cause another chemical species to undergo oxidation.

9 B. Reducing agent is synonymous with reductant.

10 B. Conjugate redox pairs are written with oxidising agent first, followed by the reducing agent.

Deconstructed questions

11 C **12** A

13 [For a complete redox equation, one conjugate redox pair will consist of the oxidising agent and the chemical species it is reduced to, while the other will consist of the reducing agent and the chemical species it is oxidised to.[1]] [Therefore, one conjugate redox pair is $Co^{2+}_{(aq)}/Co_{(s)}$ and the other is $Al^{3+}_{(aq)}/Al_{(s)}$.[2]]

✓ ✗ I have identified what conjugate redox pairs consist of.[1]

✓ ✗ I have identified the conjugate redox pairs for the forward equation.[2]

Exam-style questions

Within lesson

14 D **15** C

16 **a** [Oxidising agents cause another chemical species to be oxidised, and are themselves reduced.[1]] [Therefore, since $O_{2(g)}$ is the oxidising agent, it will be reduced.[2]]

✓ ✗ I have described what occurs to oxidising agents.[1]

✓ ✗ I have identified the chemical species being reduced.[2]

b Oxidising agent/reducing agent

$O_{2(g)}/H_2O_{(l)}$

$HNO_{3(g)}/NH_{3(g)}$

c [Nitric acid is an oxidising agent.[1]] [This is because it is the conjugate redox partner to a reducing agent, $NH_{3(g)}$, which has lost electrons to form nitric acid.[2]]

✓ ✗ I have identified the oxidising agent.[1]

✓ ✗ I have justified my answer with reference to the conjugate redox pair.[2]

d Since this is not a redox reaction, electrons are not transferred between chemical species.

17 **a** [An oxidising agent is a chemical species that causes another to be oxidised (it is reduced itself) and a reducing agent is a chemical species that causes another to be reduced (it is oxidised itself).[1]] [Therefore, $Cu^{+}_{(aq)}$ is both the oxidising agent and reducing agent because it is reduced to form $Cu_{(s)}$ and is oxidised to form $Cu^{2+}_{(aq)}$.[2]]

✓ ✗ I have described the function of oxidising and reducing agents.[1]

✓ ✗ I have linked these definitions to the role of $Cu^{+}_{(aq)}$ in the chemical reaction.[2]

b Oxidising agent/reducing agent

$Cu^{+}_{(aq)}/Cu_{(s)}$

$Cu^{2+}_{(aq)}/Cu^{+}_{(aq)}$

c [Since $O_{2(g)}$ is a product of the reaction and is an oxidising agent, this means that $H_2O_{2(g)}$ must have been oxidised to form $O_{2(g)}$.[1]] [This is because it has to have lost electrons in order for $O_{2(g)}$ to be able to gain electrons and be an oxidising agent.[2]]

✓ ✗ I have linked an oxidising agent on the product side to a chemical species that has been oxidised.[1]

✓ ✗ I have explained this connection with reference to the transfer of electrons.[2]

d [Hydrogen peroxide is unstable because it can function as both an oxidising agent and reducing agent, effectively oxidising and reducing molecules of itself.[1]] [As a result, it decomposes readily and is very unstable as it doesn't remain in its pure form for long.[2]]

✓ ✗ I have described the characteristic of $H_2O_{2(g)}$ as both an oxidising and reducing agent.[1]

✓ ✗ I have linked my answer to the question.[2]

Multiple lessons

18 **a** [Chemical compounds which are oxidised lose electrons.[1]] [Therefore, if antioxidants are oxidised in preference to another compound, they must be able to lose electrons easily.[2]]

✓ ✗ I have defined the process of oxidation.[1]

✓ ✗ I have linked the ability to be oxidised to the ability to lose electrons.[2]

b [The student is incorrect.[1]] [This is because oxidising agents cause another chemical species to be oxidised and are, therefore, reduced themselves as they gain electrons from a chemical species that has lost electrons (i.e. been oxidised).[2]]

✓ ✗ I have commented on the accuracy of the student's statement.[1]

✓ ✗ I have explained why oxidising agents are reduced themselves.[2]

c The hydroxyl functional group, OH

d [Antioxidants are chemical species which are easily able to lose electrons and be oxidised themselves.[1]] [The hydroxyl function group, with its two lone pairs of electrons on each oxygen atom, means that there are plenty of electrons that can be lost, allowing it to readily undergo oxidation which is the loss of electrons.[2]]

I have defined the purpose of an antioxidant.[1]

I have explained how the structure of the hydroxyl functional group relates to its ability to be oxidised.[2]

e [$I_{2(s)}$ is the oxidising agent because it is reduced to $I^-_{(aq)}$ as shown by the addition of negatively charged electrons.[1]][The conjugate redox pair is $I_{2(s)}/I^-_{(aq)}$.[2]]

I have identified the oxidising agent.[1]

I have identified the conjugate redox pair.[2]

f [The reducing agent is $C_6H_8O_{6(aq)}$ because it is oxidised to form $C_6H_6O_{6(aq)}$, which is the purpose of an antioxidant.[1]][Therefore, its conjugate redox pair is $C_6H_6O_{6(aq)}/C_6H_8O_{6(aq)}$.[2]]

I have identified the reducing agent.[1]

I have identified the conjugate redox pair.[2]

19 a [A fully oxidised compound cannot undergo further oxidation which is the loss of electrons.[1]][Therefore, ubiquinone is only able to be reduced and gain electrons.[2]]

I have defined a fully oxidised compound.[1]

I have linked this to its ability to only be reduced.[2]

b [While ubiquinone has some polar covalent bonds present due to the presence of electronegative oxygen, it has a very long non-polar carbon chain.[1]][Since the majority of the molecule is non-polar, it is not soluble in water and is instead more soluble in non-polar compounds.[2]]

I have described the overall polarity of the ubiquinone molecule.[1]

I have linked the molecule's polarity to its solubility.[2]

c [A lack of Coenzyme Q, an electron carrying molecule, means that electrons are carried in the body by free radicals.[1]][As a result, an increased concentration of free radicals damages the body and can therefore be life-threatening.[2]]

I have linked the inability to transport electrons to the formation of free radicals.[1]

I have linked an increase in concentration of free radicals to increased damage in the body.[2]

d [Cytochrome C carries electrons which are ultimately donated to $O_{2(g)}$, causing it to be reduced.[1]][Therefore, Cytochrome C is a reducing agent because it causes the reduction of $O_{2(g)}$.[2]]

I have explained the effect of Cytochrome C.[1]

I have identified the function of a reducing agent.[2]

Key science skills

20 a [Saxon has gathered qualitative data.[1]][This is because the data relates to experimental observations involving colour changes, not numerical information.[2]]

I have identified whether Saxon has gathered quantitative or qualitative data.[1]

I have justified my answer with reference to the definition of qualitative data.[2]

b Since $Ce^{4+}_{(aq)}$ is yellow and $Ce^{3+}_{(aq)}$ is colourless, the colour change indicates that $Ce^{3+}_{(aq)}$ is being produced.

c $Zn_{(s)}$

d [The colour change observation could be based on a certain spectrum of colours.[1]][This would ensure that when the experiment is repeated, the experimenters are all comparing the colour change to a standard set of colours.[2]]

I have identified one method to improve the reliability of the experiment.[1]

I have justified why it would improve the reliability of the experiment.[2]

9B Redox equations

Theory review questions

1 A. The transfer of electrons during a redox reaction will always result in changes in oxidation numbers.

2 A. The sum of all the oxidation numbers in neutral compounds will be equal to zero.

3 B. In most compounds, hydrogen and oxygen will have oxidation numbers of +1 and −2 respectively.

4 B. The sum of the oxidation numbers in NO_3^- will be equal to −1.

5 A. The sum of the oxidation numbers in a polyatomic ion will be equal to its charge, −3. The four oxygen atoms contribute 4 × (−2) = −8 to the oxidation number sum and nitrogen will have an oxidation number of +5.

6 A. An increase in oxidation number is due to the loss of negatively charged electrons, which occurs during oxidation.

7 B. Both overall redox equations and half-equations must be balanced with respect to both charge and the number of each element on each side of the equation.

8 B. Overall redox equations will not explicitly show the movement of electrons, but will involve the transfer of electrons.

9 A. The imbalance of hydrogen atoms from the addition of water must be balanced with $H^+_{(aq)}$ ions added to the reactants' side of the half-equation.

10 B. Redox half-equations must be balanced with respect to both charge and the number of each element on each side of the equation.

Deconstructed questions

11 B **12** A

13 (1) Oxidation: $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$

(2) Reduction: $NO_3^-{}_{(aq)} + 4H^+_{(aq)} + 3e^- \rightarrow NO_{(g)} + 2H_2O_{(l)}$

Overall: 3 × (1) + 2 × (2)

Overall: $3Cu_{(s)} + 2NO_3^-{}_{(aq)} + 8H^+_{(aq)} \rightarrow 3Cu^{2+}_{(aq)} + 2NO_{(g)} + 4H_2O_{(l)}$

Exam-style questions

Within lesson

14 C **15** D

16 **a** **i** Reduction: $Br_{2(l)} + 2e^- \rightarrow 2Br^-_{(aq)}$

ii Reduction: $Cr^{3+}_{(aq)} + 3e^- \rightarrow Cr_{(s)}$

iii Oxidation: $Ag_{(s)} \rightarrow Ag^+_{(aq)} + e^-$

iv Oxidation: $2I^-_{(aq)} \rightarrow I_{2(s)} + 2e^-$

v Reduction: $Sn^{4+}_{(aq)} + 2e^- \rightarrow Sn^{2+}_{(aq)}$

vi Oxidation: $2H_2O_{(l)} \rightarrow H_2O_{2(aq)} + 2H^+_{(aq)} + 2e^-$

b **i** Oxidation: $PO_3^{3-}_{(aq)} + H_2O_{(l)} \rightarrow PO_4^{3-}_{(aq)} + 2H^+_{(aq)} + 2e^-$

ii Oxidation: $Mn_2O_{3(aq)} + H_2O_{(l)} \rightarrow 2MnO_{2(s)} + 2H^+_{(aq)} + 2e^-$

iii Reduction: $2NO_3^-_{(aq)} + 4H^+_{(aq)} + 2e^- \rightarrow N_2O_{4(g)} + 2H_2O_{(l)}$

iv Oxidation: $Pb^{2+}_{(aq)} + 2H_2O_{(l)} \rightarrow PbO_{2(s)} + 4H^+_{(aq)} + 2e^-$

v Reduction: $ReO_4^-_{(aq)} + 8H^+_{(aq)} + 7e^- \rightarrow Re_{(s)} + 4H_2O_{(l)}$

vi Oxidation: $NH_{3(aq)} + 2H_2O_{(l)} \rightarrow NO_{2(g)} + 7H^+_{(aq)} + 7e^-$

c **i** $Br_{2(l)} + 2I^-_{(aq)} \rightarrow 2Br^-_{(aq)} + I_{2(s)}$

ii $Cr^{3+}_{(aq)} + 3Ag_{(s)} \rightarrow Cr_{(s)} + 3Ag^+_{(aq)}$

iii No reaction

iv $2NO_3^-_{(aq)} + 2H^+_{(aq)} \rightarrow H_2O_{2(aq)} + N_2O_{4(g)}$

v $7Mn_2O_{3(aq)} + 2ReO_4^-_{(aq)} + 2H^+_{(aq)} \rightarrow 14MnO_{2(s)} + 2Re_{(s)} + H_2O_{(l)}$

vi $Sn^{4+}_{(aq)} + Pb^{2+}_{(aq)} + 2H_2O_{(l)} \rightarrow Sn^{2+}_{(aq)} + PbO_{2(s)} + 4H^+_{(aq)}$

d **i** Half-equations:

$Cr_2O_7^{2-}_{(aq)} + 14H^+_{(aq)} + 6e^- \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

Overall equation:

$Cr_2O_7^{2-}_{(aq)} + 14H^+_{(aq)} + 3Zn_{(s)} \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)} + 3Zn^{2+}_{(aq)}$

ii Half-equations:

$I_{2(s)} + 6H_2O_{(l)} \rightarrow 2IO_3^-_{(aq)} + 12H^+_{(aq)} + 10e^-$

$PbO_{2(s)} + 4H^+_{(aq)} + 2e^- \rightarrow Pb^{2+}_{(aq)} + 2H_2O_{(l)}$

Overall equation:

$I_{2(s)} + 5PbO_{2(s)} + 8H^+_{(aq)} \rightarrow 2IO_3^-_{(aq)} + 5Pb^{2+}_{(aq)} + 4H_2O_{(l)}$

iii Half-equations:

$As_{(s)} + 3H^+_{(aq)} + 3e^- \rightarrow AsH_{3(s)}$

$As_{(s)} + 4H_2O_{(l)} \rightarrow H_2AsO_4^-_{(aq)} + 6H^+_{(aq)} + 5e^-$

Overall equation:

$8As_{(s)} + 12H_2O_{(l)} \rightarrow 3H_2AsO_4^-_{(aq)} + 3H^+_{(aq)} + 5AsH_{3(s)}$

17 **a** [S has an oxidation number of −2 in $ZnS_{(s)}$ and an oxidation number of +4 in $SO_{2(g)}$.[1]][This increase in oxidation number is due to the loss of negatively charged electrons, meaning that $ZnS_{(s)}$ must have undergone oxidation.[2]]

✓ ✗ I have identified the oxidation number of sulfur in $ZnS_{(s)}$ and $SO_{2(g)}$.[1]

✓ ✗ I have linked the change in oxidation number to whether $ZnS_{(s)}$ has undergone reduction or oxidation.[2]

b $S^{2-}_{(s)} + 2H_2O_{(l)} \rightarrow SO_{2(g)} + 4H^+_{(aq)} + 6e^-$

c $O_{2(g)} + 4e^- \rightarrow 2O^{2-}_{(s)}$

d One mole of $ZnO_{(s)}$ will be produced.

Multiple lessons

18 **a** [Fe in $FeCl_{3(s)}$ has an oxidation number of +3 and an oxidation number of +2 in $FeCl_{2(s)}$.[1]][This means that $FeCl_{3(s)}$ must be the oxidising agent as it has undergone reduction, shown by the decrease in oxidation number due to the gain of negatively charged electrons.[2]][S in $H_2S_{(g)}$ has an oxidation number of −2 and an oxidation number of 0 in $S_{(s)}$.[3]][This means that $H_2S_{(g)}$ must be the reducing agent as it has undergone oxidation, shown by the increase in oxidation number due to the loss of negatively charged electrons.[4]]

✓ ✗ I have described the change in oxidation number for Fe.[1]

✓ ✗ I have identified the oxidising agent in this reaction, with reference to the change in oxidation number.[2]

✓ ✗ I have described the change in oxidation number for S.[3]

✓ ✗ I have identified the reducing agent in this reaction, with reference to the change in oxidation number.[4]

b [N in $N_{2(g)}$ has an oxidation number of 0 and an oxidation number of −3 in $Mg_3N_{2(s)}$.[1]][This means that $N_{2(g)}$ must be the oxidising agent as it has undergone reduction, shown by the decrease in oxidation number due to the gain of negatively charged electrons.[2]][Mg in $Mg_{(s)}$ has an oxidation number of 0 and an oxidation number of +2 in $Mg_3N_{2(s)}$.[3]][This means that $Mg_{(s)}$ must be the reducing agent as it has undergone oxidation, shown by the increase in oxidation number due to the loss of negatively charged electrons.[4]]

✓ ✗ I have described the change in oxidation number for N.[1]

✓ ✗ I have identified the oxidising agent in this reaction, with reference to the change in oxidation number.[2]

✓ ✗ I have described the change in oxidation number for Mg.[3]

✓ ✗ I have identified the reducing agent in this reaction, with reference to the change in oxidation number.[4]

c Part a:

Oxidation: $H_2S_{(g)} \rightarrow S_{(s)} + 2H^+_{(aq)} + 2e^-$

Reduction: $Fe^{3+}_{(s)} + e^- \rightarrow Fe^{2+}_{(s)}$

Part b:

Oxidation: $Mg_{(s)} \rightarrow Mg^{2+}_{(s)} + 2e^-$

Reduction: $N_{2(g)} + 6e^- \rightarrow 2N^{3-}_{(s)}$

d Oxidising agent/conjugate reducing agent

Part a: $Fe^{3+}_{(s)}/Fe^{2+}_{(s)}$ and $S_{(s)}/H_2S_{(g)}$

Part b: $Mg^{2+}_{(s)}/Mg_{(s)}$ and $N_{2(g)}/N^{3-}_{(s)}$

19 **a** $\overset{+1}{12H^+}_{(aq)} + \overset{+5\ -2}{2ClO_3^-}_{(aq)} + 10e^- \rightarrow \overset{0}{Cl}_{2(g)} + \overset{+1\ -2}{6H_2O}_{(l)}$

b $\overset{+2\ -2}{NO}_{(g)} + \overset{+1\ -2}{2H_2O}_{(l)} \rightarrow \overset{+5\ -2}{NO_3^-}_{(aq)} + \overset{+1}{4H^+}_{(aq)} + 3e^-$

c $10NO_{(g)} + 2H_2O_{(l)} + 6ClO_3^-_{(aq)} \rightarrow 10NO_3^-_{(aq)} + 4H^+_{(aq)} + 3Cl_{2(g)}$

d [In the balanced overall redox reaction, there is a net production of $H^+_{(aq)}$ ions.[1]][This increase in the concentration of $H^+_{(aq)}$ ions will result in a decrease in pH.[2]]

✓ ✗ I have described the effect of the redox reaction on the production of $H^+_{(aq)}$ ions.[1]

✓ ✗ I have explained how this will affect the pH of the solution.[2]

Key science skills

20 **a** Colour of the solution

b 1: +7, 2: +6, 3: +5, 4: +4, 5: +4

c $MnO_4^{2-}{}_{(aq)} + e^- \rightarrow MnO_4^{3-}{}_{(aq)}$

d [As the experiment proceeds, a higher amount of $MnO_4^-{}_{(aq)}$ is reduced to $MnO_{2(s)}$ by the sugars in the lollipop.[1]][As a low amount of $MnO_{2(s)}$ leads to a yellow coloured solution, the orange colour could be a result of this increased amount of $MnO_{2(s)}$ present in the beaker.[2]]

✓ ✗ I have described the relationship between the experimental procedure and the amount of $MnO_{2(s)}$ present in the beaker.[1]

✓ ✗ I have suggested a reason for the observed change in the solution's colour.[2]

9C Metal displacement reactions

Theory review questions

1 A. Highly reactive metals readily undergo oxidation, making them strong reducing agents.

2 B. $Au^+_{(aq)}$ will gain an electron, undergoing reduction to form $Au_{(s)}$, a weak reducing agent. This makes $Au^+_{(aq)}$ a strong oxidising agent.

3 A. $Al_{(s)}$ is lower on the reactivity series than $Cu_{(s)}$, making it a stronger reducing agent and a more reactive metal.

4 A. Highly reactive metals readily undergo oxidation, donating electrons to the cation in solution to form their own cations and, in doing so, displacing the less reactive metal from solution.

5 B. For a spontaneous redox reaction, there will be a negative gradient in the line connecting the two chemical species on the reactivity series.

6 A. Metal displacement reactions are spontaneous redox reactions as they do not require any additional energy or heat to occur.

7 B. As a redox reaction, metal displacement reactions require simultaneous oxidation and reduction processes to allow the transfer of electrons.

8 B. Zinc is a more reactive metal than lead, and so will displace lead from solution.

9 B. Potassium, K, is so highly reactive that it will vigorously react with $O_{2(g)}$ from the atmosphere.

10 A. The strongest reducing agents (most reactive metals) are found on the bottom right of the reactivity series.

Deconstructed questions

11 C **12** B

13 Oxidation: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

Reduction: $Co^{2+}_{(aq)} + 2e^- \rightarrow Co_{(s)}$

Overall: $Zn_{(s)} + Co^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Co_{(s)}$

Exam-style questions

Within lesson

14 C **15** D

16 **a** $Pb^{2+}_{(aq)} + Co_{(s)} \rightarrow Pb_{(s)} + Co^{2+}_{(aq)}$

b $2Ag^+_{(aq)} + Cu_{(s)} \rightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)}$

c $3Sn^{2+}_{(aq)} + 2Al_{(s)} \rightarrow 3Sn_{(s)} + 2Al^{3+}_{(aq)}$

d No reaction

e $Mg_{(s)} + 2Au^+_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Au_{(s)}$

f No reaction

17 D

Multiple lessons

18 **a** (1) $[NO_3^-] = 2 \times [Cu(NO_3)_2] = 2 \times 1.2 = 2.4$ M

(2) $[NO_3^-] = 3 \times [Cr(NO_3)_3] = 3 \times 0.90 = 2.7$ M

Beaker 2 has a higher concentration of $NO_3^-{}_{(aq)}$ ions.

b (1) $n(NO_3^-) = 2.4 \times 0.750 = 1.8$ mol

(2) $n(NO_3^-) = 2.7 \times 0.500 = 1.4$ mol

Beaker 1 has a higher number of mol of $NO_3^-{}_{(aq)}$ ions.

c [$Ni_{(s)}$ is a stronger reducing agent than $Cu_{(s)}$, but not $Cr_{(s)}$.[1]][This means that $Ni_{(s)}$ will displace Cu^{2+} from solution, but Cr^{3+}.[2]][Therefore, the metal displacement reaction: $Ni_{(s)} + Cu^{2+}_{(aq)} \rightarrow Ni^{2+}_{(aq)} + Cu_{(s)}$ to produce $Ni^{2+}_{(aq)}$ ions will only occur in beaker 1.[3]][Hence, only beaker 1 will experience an increase in $Ni^{2+}_{(aq)}$ concentration.[4]]

✓ ✗ I have described the relative reactivity of Ni, Cu and Cr.[1]

✓ ✗ I have explained how this will affect the experimental results.[2]

✓ ✗ I have included a balanced overall redox equation in my answer.[3]

✓ ✗ I have explained how this will affect the $Ni^{2+}_{(aq)}$ concentration in each beaker.[4]

d There are no reactions occurring in either beaker that will significantly contribute to the loss or gain of $H^+_{(aq)}$ and $OH^-_{(aq)}$ ions. Therefore, the pH will not be significantly affected in either beaker.

19 **a** Oxidation: $Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2e^-$
Reduction: $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$

b $Sn_{(s)} + 2Ag^+_{(aq)} \rightarrow Sn^{2+}_{(aq)} + 2Ag_{(s)}$

c **i** During the metal displacement reaction, $Sn_{(s)}$ is lost from the tin figurine to form $Sn^{2+}_{(aq)}$ ions in the solution, but $Ag_{(s)}$ is also gained as a silver deposit from $Ag^+_{(aq)}$ ions forms on the figurine as well. This results in the observed change in mass.

ii $n(Sn_{(s)}\text{ lost}) = \frac{1}{2} \times n(Ag_{(s)}\text{ gained})$

$\Delta m(\text{figurine}) = [m(Ag_{(s)}\text{ gained})] - [m(Sn_{(s)}\text{ lost})]$

$+ 37.4 = [n(Ag_{(s)}\text{ gained}) \times M(Ag)] - [n(Sn_{(s)}\text{ lost}) \times M(Sn)]$

$+ 37.4 = [n(Ag_{(s)}\text{ gained}) \times 107.9] - [\frac{1}{2} \times n(Ag_{(s)}\text{ gained}) \times 118.7]$

$+ 37.4 = 48.55 \times n(Ag_{(s)}\text{ gained})$

$n(Ag_{(s)}\text{ gained}) = 0.770$ mol

$m(Ag_{(s)}\text{ gained}) = 0.770 \times 107.9 = 83.1$ g

Key science skills

20 **a** $4Rn_{(s)} + 5Kt^{4+}_{(aq)} \rightarrow 4Rn^{5+}_{(aq)} + 5Kt_{(s)}$

b [In experiment 1, $Rn_{(s)}$ displaces $Kt^{4+}_{(aq)}$ from the solution, shown by the green metallic coating and the solution's change in colour. This means that $Rn_{(s)}$ is a stronger reducing agent than $Kt_{(s)}$, making it a more reactive metal.[1]][In experiment 2, $Sk_{(s)}$ cannot displace $Rn^{5+}_{(aq)}$ from its solution as there is no reaction. This means that $Rn_{(s)}$ is also a stronger reducing agent than $Sk_{(s)}$, making it a more reactive metal.[2]][In experiment 3, $Sk_{(s)}$ displaces $Kt^{4+}_{(aq)}$ from the solution, shown by the green metallic coating and the solution's change in colour. This means that $Sk_{(s)}$ is a stronger reducing agent than $Kt_{(s)}$, making it a more reactive metal.[3]][Therefore, Sparkonium is more reactive than Kostrethium, but less reactive than Roanorium.[4]][The order from least reactive to most reactive: Kt, Sk, Rn.[5]]

- I have explained the results of experiment 1 with reference to any observations.[1]
- I have explained the results of experiment 2 with reference to any observations.[2]
- I have explained the results of experiment 3 with reference to any observations.[3]
- I have linked these results to each metal's relative reactivity.[4]
- I have answered the question by ranking the metals by their reactivity.[5]

9D Redox chemistry in the body

Theory review questions

1 B. Glucose is broken down in cells through cellular respiration.

2 A. Glucose is converted to CO_2, H_2O and energy in the form of ATP^{3-}.

3 B. Reduction can be defined as the gain of hydrogen and/or loss of oxygen atoms.

4 A. FAD can accept two protons when it is reduced to form $FADH_2$.

5 A. Reduction can be defined as the gain of hydrogen by a molecule.

6 A. NAD^+ and FAD are first reduced to NADH and $FADH_2$, which are then oxidised to release their electrons at the electron transport chain.

7 B. $NAD^+_{(aq)}$ can accept one $H^+_{(aq)}$ ion, becoming reduced to form $NADH_{(aq)}$.

8 B. At Complex II of the ETC, $FADH_2$ is oxidised (losing hydrogen atoms) to form FAD.

9 B. At Complex IV of the ETC, oxygen acts as the terminal electron acceptor, becoming reduced to form water, $H_2O_{(l)}$.

10 A. Disrupting the ETC leads to a build-up of protons, H^+, and low levels of energy in the form of ATP^{3-} in the body's cells and bloodstream.

Deconstructed questions

11 C **12** B

13 [Cyanide poisoning interferes with cellular respiration by inhibiting Complex IV of the ETC, which is responsible for the majority of cell energy production.[1]][This leads to a decrease in the amount of cellular energy in the form of ATP^{3-} that is present within cells.[2]][Therefore, the patient will be more likely to experience muscle fatigue due to low energy levels, rather than a sudden burst of energy.[3]]

- I have described the effect of cyanide poisoning on the ETC.[1]
- I have explained how this will affect ATP^{3-} levels in the body's cells.[2]
- I have linked my answer to the symptoms presented in the question.[3]

Exam-style questions

Within lesson

14 D

15 **a** $O_{2(g)} + 4H^+_{(aq)} + 4Fe^{2+}_{(aq)} \rightarrow 2H_2O_{(l)} + 4Fe^{3+}_{(aq)}$

b [The reduction of Fe^{3+} to Fe^{2+} in a cytochrome *c* molecule produces one electron, but the reduction of $O_{2(g)}$ requires four electrons to produce two water molecules.[1]][Therefore, four cytochrome *c* molecules are required as they each contain only one iron ion.[2]]

- I have explained the stoichiometry of the redox reaction that occurs at Complex IV.[1]
- I have used this explanation to answer the question.[2]

c [Without a functional Complex IV, the Fe^{2+} ions in cytochrome *c* will be unable to undergo oxidation to Fe^{3+}.[1]][Therefore, a patient suffering from Complex IV deficiency will have more cytochrome *c* molecules in the Fe^{2+} form.[2]]

- I have explained the effect of Complex IV deficiency on its key redox reaction.[1]
- I have used this explanation to answer the question.[2]

16 D

17 **a** Oxidation: $NADH \rightarrow NAD^+ + H^+ + 2e^-$

Reduction: $Q + 2H^+ + 2e^- \rightarrow QH_2$

Overall: $NADH + Q + H^+ \rightarrow NAD^+ + QH_2$

b Oxidation: $FADH_2 \rightarrow FAD + 2H^+ + 2e^-$

Reduction: $Q + 2H^+ + 2e^- \rightarrow QH_2$

Overall: $FADH_2 + Q \rightarrow FAD + QH_2$

c [Coenzyme Q deficiency will lead to a decrease in the reactions occurring at Complexes I and II as they require ubiquinone (Q).[1]][This means there will be a buildup of NADH in the cell as it is no longer consumed at Complex I.[2]][As the protons driven into the IMS by the reactions of Complexes I and II that are necessary for ATP^{3-} synthesis will also be reduced, there will be decreased levels of ATP^{3-} in the cell.[3]]

- I have described the effect on the reactions of Complexes I and II.[1]
- I have described the effect on NADH levels in the cell.[2]
- I have described the effect on ATP^{3-} levels in the cell.[3]

Multiple lessons

18 **a** $C_6H_{12}O_6$: 0 CO_2: +4

[Glucose has undergone oxidation[1]][as there was an increase in oxidation number caused by the loss of negatively charged electrons.[2]]

✓ ✗ I have identified the type of reaction that glucose has undergone.[1]

✓ ✗ I have used the change in oxidation number to explain my reasoning.[2]

b O_2: 0 H_2O: −2

[Oxygen has undergone reduction[1]][as there was a decrease in oxidation number caused by the gain of negatively charged electrons.[2]]

✓ ✗ I have identified the type of reaction that oxygen has undergone.[1]

✓ ✗ I have used the change in oxidation number to explain my reasoning.[2]

c $m(\text{GlucaBeans}) = 5 \times 1.12 = 5.60$ g

$m(C_6H_{12}O_6) = 0.90 \times 5.60 = 5.04$ g

$$n(C_6H_{12}O_6) = \frac{5.04}{(6 \times 12.0) + (12 \times 1.0) + (6 \times 16.0)} = 0.028 \text{ mol}$$

$$n(ATP^{3-}) = \frac{32}{1} \times 0.028 = 0.896 \text{ mol}$$

[If the GlucaBeans were 90% glucose, we would expect 0.896 mol of ATP^{3-} produced.[1]][However, as there was only 0.58 mol produced, GlucaBean's claim is likely to be false.[2]]

✓ ✗ I have included calculations in my answer.

✓ ✗ I have identified the expected amount of ATP^{3-}.[1]

✓ ✗ I have commented on the accuracy of the claim.[2]

19 a The conversion of NADH to NAD^+ is a half-reaction as it is only an oxidation reaction, not a full redox reaction.

b -OH: hydroxyl -COO^-: carboxyl (deprotonated)

c This suggests that the concentration of $H^+_{(aq)}$ ions in the IMS is 10 times higher than the concentration of $H^+_{(aq)}$ in the mitochondrial matrix.

Key science skills

20 a Quantitative

b B

c [The strongest inhibitor of *E. Chemistridium* growth will result in the lowest number of microbes in solution, observed as the lowest turbidity.[1]][This means that hydrazoic acid, HN_3, will be the most effective inhibitor as it resulted in the lowest observed turbidity of 32%.[2]]

✓ ✗ I have explained the relationship between inhibitor strength and observed turbidity.[1]

✓ ✗ I have identified the most effective inhibitor with reference to the experimental results.[2]

Chapter 9 review

Multiple choice questions

1 B. The oxidising agent undergoes reduction, decreasing in oxidation number. The reducing agent undergoes oxidation, increasing in oxidation number.

2 C. This is an oxidation half-reaction as the oxidation number of nitrogen increases from −3 to +4 due to the loss of negatively charged electrons.

3 D. An increase in oxidation number is due to an oxidation process occurring.

4 D. Four $Al_{(s)}$ atoms must each lose three electrons to form $Al^{3+}_{(aq)}$ ions, giving a total of twelve electrons that are transferred in this redox reaction.

5 A. Reducing agents are oxidised to form their conjugate oxidising agents. $PbO_{2(s)}$ is reduced to $Pb^{2+}_{(aq)}$ and this reaction must be balanced with respect to both charge and number of each element.

6 B. The solid of a more reactive metal (a stronger reducing agent and lower on the reactivity series) will displace a less reactive (a weaker reducing agent) metal's cations from solution.

7 B. Whether a metal displacement reaction will occur depends on the oxidising and reducing agent strengths of the species involved. $K_{(s)}$ is highly reactive so will react vigorously even with water.

8 A. $Q + 2H^+ + 2e^- \rightarrow QH_2$ is a reduction process. This is shown by the gain of electrons or the gain of hydrogen atoms by ubiquinone, Q, molecules.

9 D. Half-equations are balanced according to the KOHES sequence of steps.

10 C. The number of mol of $Fe^{2+}_{(aq)}$ in the solution will be proportional to the mass of $Ag_{(s)}$ deposited on the iron screw.

Short answer questions

11 a Carbon: $H_2C_2O_{4(aq)}$ (+3), $CO_{2(g)}$ (+4)

Manganese: $MnO_4^-{}_{(aq)}$ (+7), $Mn^{2+}_{(aq)}$ (+2)

b [$MnO_4^-{}_{(aq)}$ is reduced.[1]][This is because the oxidation number of Mn decreases from +7 to +2 due to the gain of negatively charged electrons.[2]][$H_2C_2O_{4(aq)}$, however, is oxidised.[3]][This is because the oxidation number of C increases from +3 to +4 due to the loss of negatively charged electrons.[4]]

✓ ✗ I have identified the species undergoing reduction.[1]

✓ ✗ I have explained my answer with respect to oxidation numbers.[2]

✓ ✗ I have identified the species undergoing oxidation.[3]

✓ ✗ I have explained my answer with respect to oxidation numbers.[4]

c Oxidation: $H_2C_2O_{4(aq)} \rightarrow 2CO_{2(g)} + 2H^+_{(aq)} + 2e^-$

Reduction: $MnO_4^-{}_{(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$

d $5H_2C_2O_{4(aq)} + 2MnO_4^-{}_{(aq)} + 6H^+_{(aq)} \rightarrow 10CO_{2(g)} + 2Mn^{2+}_{(aq)} + 8H_2O_{(l)}$

12 a Metal displacement reaction

Spontaneous redox reaction

b Oxidation: $V_{(s)} \rightarrow V^{2+}_{(aq)} + 2e^-$

Reduction: $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$

Overall equation: $V_{(s)} + Fe^{2+}_{(aq)} \rightarrow V^{2+}_{(aq)} + Fe_{(s)}$

c Oxidising agent/Reducing agent

$Fe^{2+}_{(aq)}/Fe_{(s)}$

$V^{2+}_{(aq)}/V_{(s)}$

d [$V_{(s)}$ displaced $Fe^{2+}_{(aq)}$ cations from solution but $Fe_{(s)}$ was unable to displace $V^{2+}_{(aq)}$ ions from solution, as shown by the vigorous reaction in Experiment 2, but the lack of a reaction in Experiment 1.[1]][This means that $V_{(s)}$ is a more reactive metal than $Fe_{(s)}$, and therefore a stronger reducing agent.[2]]

I have used experimental observations to justify my answer.[1]

I have described how these experimental results explain the relative reactivity of $V_{(s)}$ and $Fe_{(s)}$.[2]

e [From Experiment 2, we know that vanadium is a more reactive metal than $Fe_{(s)}$ and so will also be more reactive than $Cu_{(s)}$ as $Cu_{(s)}$ is less reactive than $Fe_{(s)}$.[1]][This means that $V_{(s)}$ is also a stronger reducing agent than $Cu_{(s)}$ and so will displace $Cu^{2+}_{(aq)}$ ions from a solution of $CuSO_{4(aq)}$. Hence, a reaction will be observed.[2]]

I have determined the relative reactivity of $V_{(s)}$ and $Cu_{(s)}$.[1]

I have explained how this will affect whether a reaction will occur between $V_{(s)}$ and $Cu^{2+}_{(aq)}$ ions.[2]

13 a [Succinate has lost two hydrogen atoms to form fumarate (oxidation) and FAD has gained two hydrogen atoms to form $FADH_2$ (reduction).[1]][This suggests that a redox reaction has occurred.[2]]

I have described the reaction involving FAD and succinate with reference to its hydrogen atoms.[1]

I have identified the type of reaction that has occured.[2]

b Oxidation: $C_4H_4O_4^{2-}{}_{(aq)} \rightarrow C_4H_2O_4^{2-}{}_{(aq)} + 2H^+_{(aq)} + 2e^-$

Reduction: $FAD_{(aq)} + 2H^+_{(aq)} + 2e^- \rightarrow FADH_{2(aq)}$

c $C_4H_4O_4^{2-}{}_{(aq)} + FAD_{(aq)} \rightarrow C_4H_2O_4^{2-}{}_{(aq)} + FADH_{2(aq)}$

d [Succinate, $C_4H_4O_4^{2-}{}_{(aq)}$, is the reducing agent in this reaction.[1]][This is because it causes the reduction of $FAD_{(aq)}$ to $FADH_{2(aq)}$ and undergoes oxidation itself to form fumarate, $C_4H_2O_4^{2-}{}_{(aq)}$.[2]]

I have identified the reducing agent.[1]

I have explained why this chemical species is defined as the reducing agent.[2]

e $FADH_2 + Q \rightarrow FAD + QH_2$

14 a $\overset{+5\ -2}{NO_3^-{}_{(aq)}} + \overset{+1\ +1\ -2}{HBrO_{(aq)}} \rightarrow \overset{+4\ -2}{N_2O_{4(g)}} + \overset{+5\ -2}{BrO_3^-{}_{(aq)}}$

b $4NO_3^-{}_{(aq)} + 3H^+_{(aq)} + HBrO_{(aq)} \rightarrow 2N_2O_{4(g)} + 2H_2O_{(l)} + BrO_3^-{}_{(aq)}$

c $n(N_2O_{4(g)}) = 2.78$ mol

$n(e^-) = \frac{4}{2} \times n(N_2O_{4(g)}) = 2 \times 2.78 = 5.56$ mol

$N(e^-) = 5.56 \times 6.02 \times 10^{23} = 3.35 \times 10^{24}$ electrons

Key science skills questions

15 a Surface area: quantitative

Absorption: quantitative

Description: qualitative

b [The white sheet of paper allows all colour descriptions and absorption readings to be taken with a similar background.[1]][This allows some variability in these readings to be controlled.[2]]

I have described the purpose of the white sheet of paper for data collection.[1]

I have explained how this influences the variability of the data obtained.[2]

c $2Ag^+_{(aq)} + Co_{(s)} \rightarrow 2Ag_{(s)} + Co^{2+}_{(aq)}$

d As the surface area of the cobalt wire increases, the rate at which $Co_{(s)}$ is oxidised to $Co^{2+}_{(aq)}$ ions in the silver nitrate solution will increase.

e [No, Meredith could not use this observation to estimate the surface area of the cobalt wire in her phone.[1]][This is because the 'ripe tomato' red description of the $Co^{2+}_{(aq)}$ solution is purely qualitative and so cannot be used to make inferences about quantitative measurements such as the wire's surface area.[2]]

I have stated whether Meredith could use this observation to estimate the wire's surface area.[1]

I have explained why this is the case with reference to qualitative and quantitative data.[2]

REVIEW

10A Water on Earth

Theory review questions

1 B. Water moves through a cycle.

2 B. The human body is not efficient at removing salt.

3 A. The water that is available to houses comes from catchments.

4 B. Most freshwater needs to be treated before consumption.

5 B. The vast majority of water on Earth is seawater, which is not safe for human consumption.

Deconstructed questions

6 D **7** A

8 [Trees and plants play an important role in the water cycle as the roots collect water from the ground and release it into the atmosphere through transpiration.[1]][The water condenses, forming clouds, which eventually results in rain.[2]][As a result, clearing plants and trees would reduce the amount of water transpiring, leaving less water available to support rainfall.[3]]

- ✓ ✗ I have described the role of trees and plants in the water cycle.[1]
- ✓ ✗ I have identified the process following transpiration.[2]
- ✓ ✗ I have described the impact of trees and plants on rainfall.[3]

Exam-style questions

Within lesson

9 a Level of salt

b [Seawater and brackish water both contain too much salt for human consumption,[1]][however freshwater can in some cases be safe.[2]][Freshwater defines a wide array of sources of water with varying concentrations of salt, and only freshwater with very low salt concentrations is considered potable.[3]]

- ✓ ✗ I have identified that seawater and brackish water are unsafe for human consumption.[1]
- ✓ ✗ I have identified that freshwater is sometimes safe for human consumption.[2]
- ✓ ✗ I have explained why not all freshwater is safe for human consumption.[3]

c i Public access is limited to prevent contamination of water supplies.

ii [Water travels from a water source through treatment plants and water mains to smaller service reservoirs.[1]][Smaller pipes deliver water directly to people's houses from service reservoirs.[2]]

- ✓ ✗ I have identified the use of water mains in water transport.[1]
- ✓ ✗ I have identified the use of smaller pipes in water transport.[2]

Multiple lessons

10 a [Water contains polar covalent bonds, and is itself an overall polar compound.[1]][As they are made of charged ions, ionic compounds dissociate when in water.[2]][As a result, ionic compounds in water would exist as ions and therefore are not detectable in water.[3]][Consequently, we are unable to 'see' any sodium chloride.[4]]

- ✓ ✗ I have described the polar nature of water.[1]
- ✓ ✗ I have described the interaction between water and ionic compounds.[2]
- ✓ ✗ I have identified how ionic compounds exist in water.[3]
- ✓ ✗ I have have linked my answer to the question.[4]

b $m(\text{NaCl})_{250\text{ mL}} = 12 \times \frac{250}{1000} = 3.0\text{ g}$

$n(\text{NaCl})_{250\text{ mL}} = \frac{3.0}{(23.0 + 35.5)} = 0.0513\text{ mol}$

c $c(\text{NaCl})_{250\text{ mL}} = \frac{0.0513}{(250 \div 1000)} = 0.205\text{ M}$

[Since the concentration of salt in the 250 mL sample of seawater is 0.205 M, it is much higher than the acceptable drinking water for humans with a concentration of 0.008 M.[1]][As a result, the sample of seawater is not safe to consume.[2]]

- ✓ ✗ I have compared the concentration of the sample and the safe drinking concentration.[1]
- ✓ ✗ I have linked my answer to the question.[2]

Key science skills

11 a The degree of knowledge of the water cycle.

b [The purpose of the experiment is to determine the understanding of the water cycle in students who had covered this information in a class.[1]][As part of the study, Segnish included both students who had and had not attended the class.[2]][This would affect the results as individuals who had not attended the class would not provide a reliable indication of the class's knowledge of the topic.[3]][This decreases the validity of the experiment and the data obtained.[4]]

- ✓ ✗ I have identified the aim of the experiment.[1]
- ✓ ✗ I have identified a key variability in the experiment.[2]
- ✓ ✗ I have explained the effect of this variability on the results.[3]
- ✓ ✗ I have linked my answer to the question.[4]

c Make sure that the people involved in the experiment have similar levels of exposure to the water cycle.

10B Water contamination

Theory review questions

1 B. Water sampling must be carried out at various locations and depths.

2 B. Water sampling is prioritised when an event that may pollute water supplies has taken place.

3 A. Agricultural and industrial activity both contribute to the chemical contamination of waterways.

4 B. Heavy metals have a high affinity for certain groups on proteins.

5 A. Organometallic compounds are organic molecules containing metals.

6 A. Stearate anions react with metal cations preventing their dirt-removal action.

7 B. Bioaccumulation explains how species lower in the food chain contain heavy metal which enters and accumulates over time in species higher in the food chain.

8 A. Persistent Organic Pollutants were greatly restricted in the Stockholm Convention due to their high toxicity.

9 B. Dioxins are toxic waste by-products from processes such as metal smelting and paper production.

10 B. Oil spills block out sunlight and are poisonous, having disastrous effects on marine ecology.

Deconstructed questions

11 C **12** D

13 [Water that does not readily produce lather from soap is known as hard water.[1]][Water sampling tests would be expected to indicate heightened levels of metal cations in the water.[2]][This is because the stearate anions interact with metal cations in the water, preventing them from carrying out their dirt-removal function.[3]][Since the presence of algae blooms in water is unrelated to the formation of hard water, it is impossible to determine what the expected results of these tests would be.[4]]

- ✓ ✗ I have identified the water type.[1]
- ✓ ✗ I have predicted the expected results of the test for metal ion contamination.[2]
- ✓ ✗ I have explained how metal cations prevents a lather from forming.[3]
- ✓ ✗ I have explained the ability to determine the results of algal contamination testing.[4]

Exam-style questions

Within lesson

14 a Answers may include at least one of:

- Arsenic
- Cadmium
- Chromium
- Copper
- Nickel
- Lead
- Mercury

b i Heavy metals have a higher density than other metals.

ii [Heavy metals are dangerous due to their high affinity for thiol groups on proteins,[1]][an interaction that occurs as a result of the high density of heavy metals.[2]][As a result of this affinity for thiol groups, these heavy metals can destabilise the stability of protein structures, compromising their function. This can be life-threatening when proteins essential to key metabolic processes are compromised.[3]]

- ✓ ✗ I have described heavy metals' high affinity for groups in proteins.[1]
- ✓ ✗ I have explained the interactions between proteins and heavy metals.[2]
- ✓ ✗ I have explained why this high affinity can be life-threatening.[3]

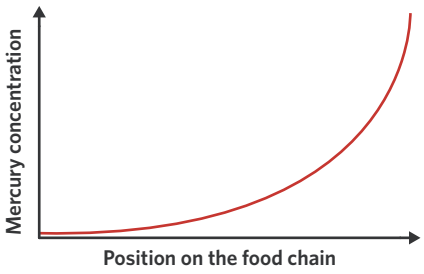
c Answers may include at least two of the following:

- Geological sources – heavy metals found in the Earth's crust
- Industrial run-off (includes mines, foundries and smelters)
- Domestic effluent
- Atmospheric heavy metals

d i [Bioaccumulation describes the phenomenon in which species higher on the food chain consume species lower in the food chain containing heavy metals.[1]][As species higher in the food chain consume more of species lower, the heavy metal found in these species lower in the food chain will gradually accumulate in the higher species due to heavy metals not biodegrading.[2]][Consequently, the species highest in the food chain will have the highest heavy metal concentration in their systems.[3]]

- ✓ ✗ I have identified the relationship in bioaccumulation regarding position on the food chain.[1]
- ✓ ✗ I have described how organisms consuming other organisms leads to them absorbing heavy metals from the others into their system.[2]
- ✓ ✗ I have concluded that the species highest on the food chain has the highest heavy metal concentration.[3]

ii



- ✓ ✗ I have included 'position on the food chain' on the x-axis.
- ✓ ✗ I have included 'heavy metal concentration' on the y-axis.
- ✓ ✗ I have drawn an exponential relationship between position on food chain and heavy metal concentration.

15 a POPs are referred to as 'forever chemicals' as they are often resistant to biodegradation.

b [As POPs are resistant to biodegradation, this means that they remain in the human body or environment for extensive periods of time.[1]][The presence of toxic molecules in the human body can lead to severe chronic medical conditions, and extended presence in the environment can cause significant damage to ecological processes.[2]]

- I have explained how POPs' resistance to biodegradation affects their interaction with the human body.[1]
- I have described some consequences of the extended presence of toxic materials.[2]

c Answers may include at least two of:

- Pesticides
- Dioxins
- Oil spills

d [A pesticide is a substance that either kills or deters pests, including insects, rodents or other forms of organisms that may be detrimental to farming practice.[1]][The use of POPs as pesticides in agriculture is dangerous as these toxic substances may find their way into agricultural run-off, entering the environment and causing both ecological and health damage.[2]]

- I have described the function of a pesticide.[1]
- I have explained the environmental and health impacts of the use use of POPs as a pesticide.[2]

e [POPs are generally non-polar in structure, which makes them remain in the human body for extensive periods of time.[1]][This is because non-polar substances are stored in non-polar fatty tissue, rather than being found in the bloodstream.[2]][As a consequence of this, rather than being excreted, POPs will remain in fatty tissue unfiltered for extensive periods of time.[3]]

- I have identified that POPs are generally non-polar in structure.[1]
- I have explained that non-polar substances are stored in fatty tissue.[2]
- I have explained that substances stored in fatty tissue are not actively excreted.[3]

16 a Crude oil is a non-polar substance.

b [The oil spill would be expected to be in the form of an oil slick covering the surface of the water, as well as much of the shoreline.[1]][This is because the oil is non-polar, and therefore does not dissolve in the water, and instead binds into a single surface slick.[2]][Oil would also be expected to be covering the shoreline and wildlife due to its ability to stick to the surface of materials it comes into contact with.[3]]

- I have described the expected appearance of the oil spill.[1]
- I have explained how crude oil's non-polarity contributes to the formation of an oil slick.[2]
- I have explained why crude oil sticks to the shoreline.[3]

c Points that may be discussed include:

- An oil slick prevents sunlight from reaching the seabed, preventing ecological processes.
- Oil can cover wildlife with oil, especially near shores where much marine life is present.
- Ingested oil is toxic for animals.
- Disrupted marine populations can interrupt the food chain.
- Oil can damage plants, damaging a food resource.
- Oil can damage fish eggs, damaging the reproduction of a species.

d [Promoting the growth of microbiological agents to break down oil spills can be effective in reducing the severity of the oil spill.[1]][Nonetheless, the rate at which these agents function is far too slow to counteract the impact of the spill entirely.[2]]

- I have identified that microbiological agents can reduce the severity of an oil spill.[1]
- I have explained that the rate of breakdown of microbiological agents.[2]

17 a Blue-green algae and E. coli

b [Chlorine is not a contaminant in this situation[1]][as it is intentionally being added as a disinfectant. Therefore, it is not polluting the water supply, and is being added at levels that are considered safe.[2]][However, if chlorine levels exceed the levels considered to be healthy by water authorities, then it becomes a contaminant.[3]]

- I have identified that chlorine is not a contaminant.[1]
- I have explained that chlorine is not polluting the water supply.[2]
- I have identified a situation in which chlorine is a contaminant.[3]

c [Saline water is water that contains high amounts of dissolved salt, such as sodium chloride.[1]][If someone were to consume saline water, the concentration of ions would be too great to filter it out effectively.[2]][The resulting accumulation of ions in the body would be detrimental to the person's health.[3]]

- I have identified the ionic composition of saline water.[1]
- I have explained that the ions would not be able to be filtered out.[2]
- I have explained the health impacts of consuming saline water.[3]

Multiple lessons

18 a [Soaps remove dirts from materials through the action of stearate ions.[1]][Sodium stearate ions in soaps dissociate into sodium cations and stearate anions, with the stearate anions able to interact with dirts for their removal.[2]][This interaction can be observed as the development of a lather or foam from soapy water.[3]]

- I have identified how soaps remove dirts from materials.[1]
- I have described the process through which stearate ions remove dirt.[2]
- I have identified that stearate interactions lead to the development of a lather.[3]

b [Stearate anions have a high solubility in water.[1]][This is because ions are able to form dipole dipole bonds with water molecules, which is strong enough to overcome the intermolecular forces between water molecules and between the sodium and stearate ions.[2]][As an ion, stearate anions are able to bind to other metallic ions to form a precipitate and is therefore able to remove contaminants in water.[3]]

- ✓ ✗ I have identified the degree of solubility of stearate ions in water.[1]
- ✓ ✗ I have explained the implications of the interaction between stearate ions and water.[2]
- ✓ ✗ I have described the function of stearate ions.[3]

c [Water with large concentrations of metallic ions is known as hard water.[1]][Soap is less effective in hard water because metallic cations in the water interact with stearate anions to form a precipitate.[2]][Consequently, the stearate is no longer active and cannot carry out its dirt removal function.[3]]

- ✓ ✗ I have identified that water with large metallic ion concentrations is hard water.[1]
- ✓ ✗ I have explained that metallic cations interact with stearate anions.[2]
- ✓ ✗ I have explained that stearate anions can no longer carry out their function having interacted with metallic cations.[3]

d [Some detergents have anions added to them so that these anions react with metallic cations in water.[1]][This prevents metallic cations from interacting with stearate anions, enabling stearate ions to carry out their dirt-removal function.[2]]

- ✓ ✗ I have identified the function of added anions.[1]
- ✓ ✗ I have explained why the function of added anions is useful.[2]

19 a Organic molecules are chemical compounds that contain carbon-hydrogen covalent bonds.

b i The movement of water between different locations and different states of matter on Earth.

ii [Freshwater, brackish water and seawater differ in the concentration of salt.[1]][The only one of these that may be safe to consume is freshwater, as this describes water with the lowest concentrations of salt.[2]][However, the term freshwater encompasses many different concentrations of salt, so only those at the lower end will be safe for human consumption.[3]]

- ✓ ✗ I have identified the difference between these types of water.[1]
- ✓ ✗ I have explained that freshwater is the only type of water that may be safe for consumption.[2]
- ✓ ✗ I have explained that not all freshwater is safe for human consumption.[3]

c i [Yes, salt can be considered a contaminant of water.[1]][As the addition of too much salt to water makes it inappropriate for human consumption and must be removed for human use, this can be considered a contaminant.[2]][Additionally, substances do not have to be introduced by humans for them to be considered contaminants, with other natural substances such as algae also being considered as contaminants.[3]]

- ✓ ✗ I have identified salt as a contaminant.[1]
- ✓ ✗ I have explained the impact of salts on water for human consumption.[2]
- ✓ ✗ I have identified the nature of the classification of contaminants.[3]

ii [No, these salts are not organic.[1]][The term organic refers to substances that possess at least one carbon-hydrogen covalent bond, and does not refer to a substance being naturally formed or not.[2]]

- ✓ ✗ I have identified that these salts are not organic.[1]
- ✓ ✗ I have described why salts are not organic.[2]

Key science skills

20 a [This is not a good representation of the water composition throughout the whole river.[1]][The composition of the water can change at different locations along the water, so gathering all samples from one location will not be fully representative.[2]][Additionally, since there were only three samples collected, this may not be enough variety in depths to obtain a full understanding of the compositional variations of the water, as there may be changes in between the depths of collection.[3]][However, this depends on the total depth of the water at the sampling location.[4]]

- ✓ ✗ I have identified that the samples collected will not be representative of the composition of the entire river.[1]
- ✓ ✗ I have explained that the composition of water can vary in different locations along the river.[2]
- ✓ ✗ I have explained why it is possible that an insufficient number of depths may have been collected.[3]
- ✓ ✗ I have identified that the number of depths of collection is dependent on the depth of the water source.[4]

b A random error.

c A systematic error.

d [Since the contamination makes each of the results different to the correct results, the results are inaccurate.[1]][However, since each sample is contaminated to the same degree by the same source, the results should all be close to each other. Thus, they would still be expected to be precise.[2]]

- ✓ ✗ I have explained how the skewed results due to contamination make the results inaccurate.[1]
- ✓ ✗ I have explained why each result being affected to the same degree makes the results still precise.[2]

e Must include three of the following:

- The container must not be reactive with the sample.
- The container must be rinsed before taking another sample.
- Water used to clean the container should be tested to ensure it has not been contaminated.
- Ensure the same chemicals are being analysed in each experiment.
- Ensure the use of the same equipment.
- Gather the required sample size.
- Ensure the same method is used for producing quantitative results of water quality.
- Ensure standardised labelling, storage and transport.

10C Analysing water using gravimetric analysis

Theory review questions

1 A. Ions conduct electricity, so electrical conductivity testing provides information about the concentration of ions in water.

2 A. Salt ions conduct electricity better than pure water.

3 B. The light intensity from a lightbulb indicates the amount of electricity passing through the circuit.

4 B. Gravimetric analysis involves precipitating ions out of solution and weighing the precipitate.

5 A. Water molecules remaining from improper drying will increase the recorded mass of a precipitate.

6 A. The product of the reaction between the analyte and the reactant is insoluble in water.

7 B. Mass-mass stoichiometry in gravimetric analysis is used to measure the mass of an analyte.

8 B. There are four NH_3 molecules reacting for every three O_2 molecules.

9 A. Mole ratio of $H_2 : O_2$ is 2 : 1, and $5 \div 3 = 1.67 < 2$, so H_2 is the limiting reactant.

10 B. Molar ratios indicate the relative numbers of moles of substances, and do not reflect their masses.

Deconstructed questions

11 C **12** A

13 [Advantages of using gravimetric analysis include that not many materials are required for chemical analysis to be carried out. Specifically, gravimetric analysis only requires an analyte and a means of drying and weighing the precipitate. It can also be used selectively for a particular ion.[1]][However, some disadvantages include the possibility of incorrectly weighing the sample due to the presence of water contaminants. Additionally, the reaction between the solute and the analyte may be slow.[2]][Advantages of using electrical conductivity testing as a form of chemical analysis include that it can be carried out quickly and easily, and the equipment used is portable.[3]][However, a disadvantage of using electrical conductivity testing is that it cannot be used to determine the concentration of a particular element, rather providing an indication of the overall ionic content of a sample.[4]]

✓ ✗	I have identified several advantages of gravimetric analysis.[1]
✓ ✗	I have identified several disadvantages of gravimetric analysis.[2]
✓ ✗	I have identified several advantages of electrical conductivity testing.[3]
✓ ✗	I have identified a disadvantage of electrical conductivity testing.[4]

Exam-style questions

Within lesson

14 a i $n(MgSO_4) : n(Ca(NO_3)_2) = 1 : 1$

$n(MgSO_4) = 0.332$ mol, $n(Ca(NO_3)_2) = 1.98$ mol

$0.332 < 1.98$

$MgSO_4$ is the limiting reactant.

ii $n(AgNO_3) : n(CaCl_2) = 2 : 1$

$n(AgNO_3) = 3.71$ mol, $n(CaCl_2) = 1.94$ mol

To use all of the $AgNO_3$ available, the $CaCl_2$ required is:

$$n(CaCl_2) = \frac{1}{2} \times 3.71 = 1.86 \text{ mol}$$

$AgNO_3$ is the limiting reactant.

iii $n(Pb(NO_3)_2) : n(NaCl) = 1 : 2$

$$n(Pb(NO_3)_2) = \frac{0.866}{207.2 + (2 \times 14.0) + (6 \times 16.0)} = 0.00261 \text{ mol},$$

$$n(NaCl) = \frac{2.34}{23.0 + 35.5} = 0.04 \text{ mol}$$

To use all the $Pb(NO_3)_2$ available, the NaCl required is:

$$n(NaCl) = \frac{2}{1} \times 0.00261 = 0.00522 \text{ mol}$$

$Pb(NO_3)_2$ is the limiting reactant.

iv $n(Al) : n(O_2) = 4 : 3$

$$n(Al) = \frac{1.45}{27.0} = 0.0537 \text{ mol}, n(O_2) = \frac{3.11}{32.0} = 0.0972 \text{ mol}$$

To use all of the Al available, the O_2 required is:

$$n(O_2) = \frac{3}{4} \times 0.0537 = 0.0403 \text{ mol}$$

Al is the limiting reactant.

b i $n(CaSO_4) : n(MgSO_4) = 1 : 1$

$n(CaSO_4) = 0.332$ mol

$m(CaSO_4) = 0.332 \times (40.1 + 32.1 + (4 \times 16.0)) = 45.2$ g

ii $n(AgNO_3) : n(AgCl) = 2 : 2$

$$n(AgCl) = \frac{2}{2} \times 3.71 = 3.71 \text{ mol}$$

$m(AgCl) = 3.71 \times (107.9 + 35.5) = 532$ g

iii $n(PbCl_2) : n(Pb(NO_3)_2) = 1 : 1$

$n(PbCl_2) = 0.00261$ mol

$m(PbCl_2) = 0.00261 \times (207.2 + (2 \times 35.5)) = 0.727$ g

iv $n(Al_2O_3) : n(Al) = 2 : 4$

$$n(Al_2O_3) = \frac{2}{4} \times 0.0537 = 0.0269 \text{ mol}$$

$m(Al_2O_3) = 0.0269 \times (2 \times 27.0) + (3 \times 16.0) = 2.74$ g

Multiple lessons

15 a Aluminium phosphate

b $n(AlPO_4) = \frac{7.78}{27.0 + 31.0 + (4 \times 16.0)} = 6.38 \times 10^{-2}$ mol

c The molar ratio is 1 : 1

d $m(Al(NO_3)_3) = (6.38 \times 10^{-2}) \times (27.0 + (14.0 \times 3) + (16.0 \times 9))$
$= 13.6$ g

16 a [Silver nitrate is required in excess.[1]][This is to ensure that all chloride ions react to form precipitate.[2]]

✓ ✗ I have identified the substance required in excess.[1]

✓ ✗ I have explained my answer based on the requirements of the reaction.[2]

b [The precipitate must have water particles removed from it before it can be weighed.[1]][This is because the precipitate may have water molecules on it when it is taken out of the solution.[2]][If still present when the precipitate is weighed, an inaccurate measurement of the mass of the sample would be obtained.[3]][The mass of the precipitate recorded would be greater than the true value, which would result in the ion concentration calculated to be higher than expected.[4]]

✓ ✗ I have identified the necessary step prior to weighing.[1]

✓ ✗ I have explained the effect of this step on the precipitate.[2]

✓ ✗ I have explained how the presence of water molecules may lead to weighing inaccuracies.[3]

✓ ✗ I have described the effect of water molecules on the results of the experiment.[4]

c i $n(AgCl) = \frac{167.4}{107.9 + 35.5} = 1.167$ mol

$n(AgCl) : n(Cl^-) = 1 : 1$

$m(Cl^-) = 1.167 \times 35.5 = 41.4$ g

ii $c(Cl^-) = \frac{1.17}{2.0} = 0.58$ M

17 a [Stearate anions are less effective in hard water as they react with cations present in the water.[1]][Stearate anions are required to be in their ionic form in order to carry out their dirt-removal function.[2]][However when they react with cations they form an insoluble and unreactive precipitate, making them less effective.[3]]

✓ ✗ I have identified why stearate anions are less effective in hard water.[1]

✓ ✗ I have explained the role of stearate ions in soaps.[2]

✓ ✗ I have explained the effect of the inability of stearate ions to function.[3]

b i [If the identity of the ions in the precipitate was unknown, then this would be an ineffective method.[1]][This is because gravimetric analysis provides quantitative data, and requires the molar mass of the precipitate which can only be determined if the identity of the constituent ions are known.[2]]

✓ ✗ I have identified the effectiveness of the method.[1]

✓ ✗ I have explained the requirements of this analytical method.[2]

ii [The efficacy of this method in determining the concentration of soap ions depends on which ions are in excess.[1]][If salt cations from the water are in excess, then the method will be effective as all of the soap ions will be present in the precipitate and measurable.[2]][However, if the soap ions are in excess, then not all of the soap ions will react and therefore this would not be an effective method of measuring the concentration of soap ions.[3]]

✓ ✗ I have identified a factor influencing the efficacy of the method.[1]

✓ ✗ I have explained the effectiveness of the method if salt ions are in excess.[2]

✓ ✗ I have explained the effectiveness of the method if soap ions are in excess.[3]

iii [If the experiment is conducted correctly, the analyte will be the limiting reactant, and the added reactant will be the excess reactant.[1]]

✓ ✗ I have identified the limiting reactant.[1]

c i $m(Mg(C_{18}H_{35}O_2)_2) = 61.0$ g

$n(Mg(C_{18}H_{35}O_2)_2) = \frac{61.0}{24.3 + 2 \times (18 \times 12.0 + 35 \times 1.0 + 2 \times 16.0)}$
$= 0.103$ mol

$n(Mg(C_{18}H_{35}O_2)_2) : n(C_{18}H_{35}O_2^-) = 1 : 2$

$n(C_{18}H_{35}O_2^-) = \frac{2}{1} \times 0.103 = 0.207$ mol

$c(C_{18}H_{35}O_2^-) = \frac{0.207}{0.500} = 0.413$ M

ii $m(C_{18}H_{35}O_2^-)$
$= 0.207 \times ((18 \times 12.0) + (35 \times 1.0) + (2 \times 16.0)) = 58.5$ g

Percentage of soap made of stearate by mass
$= \frac{58.5}{300.0} \times 100\% = 19.5\%$

18 a [Calcium is the limiting reactant.[1]][This is because if calcium is the limiting reactant, the reaction will stop when all calcium has reacted, enabling all of the calcium present to be weighed and the amount measured.[2]]

✓ ✗ I have identified the limiting reactant.[1]

✓ ✗ I have explained my answer with respect to the precipitate that must form.[2]

b i Answers may include:

- Ensure that all of the precipitate has been collected.
- Ensure that all water has been removed from the precipitate.

ii $n(Ca_3(PO_4)_2) = \frac{87.4}{3 \times 40.1 + 2 \times (31.0 + 4 \times 16.0)} = 0.282$ mol

$n(Ca^{2+}) : n(Ca_3(PO_4)_2) = 3 : 1$

$n(Ca^{2+}) = \frac{3}{1} \times 0.282 = 0.845$ mol

iii $m(Ca^{2+}) = 0.845 \times 40.1 = 33.9$ g

$$\% \text{ calcium in gypsum} = \frac{33.9}{150.0} \times 100\% = 22.6\%$$

19 a Analyte

b [It will not be possible to use octane as the analyte in this experiment.[1]] [Gravimetric analysis requires the analyte to be dissolved in a solvent so that it can be precipitated out through the addition of a reactant to the solution.[2]] [Since octane is a non-polar molecule, it will not be able to dissolve in the water. This is because octane is unable to form sufficiently strong intermolecular bonds with the water molecules in order to overcome the intermolecular forces between the water molecules.[3]]

- I have identified that octane cannot be used as the analyte in this experiment.[1]
- I have described the importance of the type of analyte used in gravimetric analysis.[2]
- I have explained the ability of octane to dissolve in water.[3]

c [A potential change that could be made would be using a non-polar solvent instead of water.[1]] [This would enable octane to dissolve in the solvent as it would not need to overcome the strong intermolecular bonds found in water.[2]]

- I have identified the nature of the solvent required.[1]
- I have explained the benefit of this change.[2]

d [It is not necessary to carry out this experiment.[1]] [The purpose of gravimetric analysis is to determine the mass of an analyte present in either a mixed sample or in a solution through precipitating out that particular analyte with a reactant and measuring the mass.[2]] [Since Vinh has proposed to use a sample of pure octane for gravimetric analysis this process is unnecessary, since by weighing the pure sample the mass of the octane can be determined.[3]]

- I have identified the necessity to perform the experiment.[1]
- I have described the purposes of gravimetric analysis.[2]
- I have explained that the pure sample of octane can be weighed.[3]

e i Potential sources of error include:

- Presence of water particles on the precipitate as a proper technique of washing, filtering, heating and cooling has not been followed.
- Some precipitate may be lost while drying with a paper towel.
- A sufficient amount of time may not have passed to form all the precipitate.

ii $MgCl_{2(aq)} + 2AgNO_{3(aq)} \rightarrow 2AgCl_{(s)} + Mg(NO_3)_{2(aq)}$

iii $$n(MgCl_2) = \frac{33.2}{24.3 + (2 \times 35.5)} = 0.348 \text{ mol}$$

$$n(AgNO_3) = \frac{108.0}{107.9 + 14.0 + (3 \times 16.0)} = 0.636 \text{ mol}$$

$n(MgCl_2) : n(AgNO_3) = 1 : 2$

To use all of the $MgCl_2$, the $AgNO_3$ required is:

$$n(AgNO_3) = \frac{2}{1} \times 0.348 = 0.696 \text{ mol}$$

Limiting reactant: $AgNO_3$

Excess reactant: $MgCl_2$

iv $n(AgCl) : n(AgNO_3) = 2 : 2$

$$n(AgCl) = \frac{2}{2} \times 0.636 = 0.636 \text{ mol}$$

$m(AgCl) = 0.636 \times (107.9 + 35.5) = 91.2$ g

Key science skills

20 a [Accuracy refers to the closeness of measurements to the true value.[1]] [Precision refers to the closeness of measurements to each other.[2]]

- I have defined the term accuracy.[1]
- I have defined the term precision.[2]

b i The results are precise, but inaccurate.

ii [The student has made a systematic error.[1]] [This is because the results deviate from the true value by a consistent amount in the one direction.[2]]

- I have identified the type of error.[1]
- I have explained why the error is a systematic error.[2]

10D Analysing water using light - part 1

Theory review questions

1 A. Sound waves are not a form of electromagnetic radiation.

2 B. Complementary colours will have opposite interactions with a substance.

3 A. Calibration curves are constructed on the basis of the direct relationship between ion concentration and absorbance to determine the concentration of ions in a solution with known absorbance.

4 B. The unit on the x-axis is the concentration which can be determined from the absorbance value.

5 B. The colour of an object is determined by the wavelength reflected or transmitted through the object to the greatest extent.

6 B. Spectrophotometers can scan a number of different electromagnetic wavelengths to choose the one with highest absorbance.

7 B. A monochromator is used to specifically select the wavelength of light to be used in UV-visible spectroscopy.

8 A. The absorbance is calculated from a series of solutions with a specific concentration of solute.

9 B. Colorimetry and UV-visible spectroscopy enable assessment of the contents in blood, such as haemoglobin, bilirubin, cholesterol and sugars.

10 A. Metal complexes enable the absorption of electromagnetic radiation by a sample.

Deconstructed questions

11 A **12** B

13 [In both colorimetry and UV-visible spectroscopy, the wavelength of maximum absorbance is used to determine the substances present; the absorbance at a particular wavelength of light is used to measure the concentration of ions in a solution.[1]] [This information is able to be obtained since specific ions strongly absorb particular wavelengths of electromagnetic radiation.[2]] [The amount of absorbance increases linearly as the concentration of ions increases, hence the concentration of the solution can be ascertained from absorbance data obtained from the colorimeter or spectrophotemeter.[3]]

- ✓ ✗ I have identified that colorimetry and UV-visible spectroscopy provide both qualitative and quantitative data.[1]
- ✓ ✗ I have explained that ions strongly absorb particular wavelengths of light.[2]
- ✓ ✗ I have explained that absorbance is directly proportional to ion concentration.[3]

Exam-style questions

Within lesson

14 a A calibration curve is a graph that visualises the relationship between the concentration of several standard solutions and their corresponding absorbance values.

b [A calibration curve is produced by preparing a series of standard solutions containing known concentrations of the solute being analysed.[1]][By measuring the absorbance of a certain wavelength of light by these standard solutions, the relationship between concentration (on the x-axis) and absorbance (on the y-axis) can be plotted on a graph to construct a calibration curve.[2]]

- ✓ ✗ I have identified the use of standard solutions in constructing a calibration curve.[1]
- ✓ ✗ I have described how a calibration curve is developed.[2]

c [Since a calibration curve represents the relationship between concentration and absorbance of a substance, it can be used to determine the concentration of a solution with a known absorbance value.[1]][The known value of absorbance can be matched up with the corresponding concentration value using the calibration curve.[2]]

- ✓ ✗ I have identified the use of a calibration curve.[1]
- ✓ ✗ I have explained how a calibration curve can be used to determine a solution's concentration.[2]

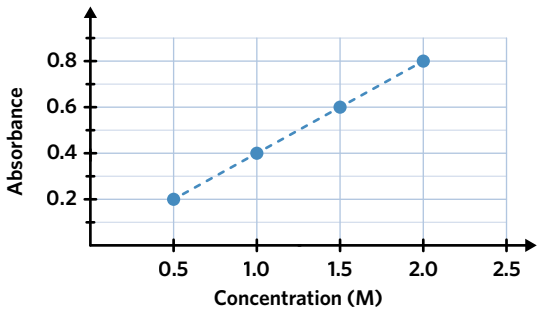
d 0.18 M

e [A solution that appears blue absorbs the complimentary colour to blue, which is orange.[1]][The copper in the solution will absorb orange as a result of the chemical properties of the copper.[2]] [This means that orange frequencies of light will need to be used when performing colorimetry.[3]]

- ✓ ✗ I have identified that orange light will be used.[1]
- ✓ ✗ I have explained why the solution absorbs orange light.[2]
- ✓ ✗ I have connected my answer to the question.[3]

15 a Absorbance in colorimetry does not have a unit.

b



- ✓ ✗ I have correctly labelled the axes.
- ✓ ✗ I have used the correct units.
- ✓ ✗ I have drawn a correctly shaped curve.

c [Ryan's claim is incorrect.[1]][This is because colorimetry measures the amount of absorbance of a particular wavelength of light, and therefore the wavelength most readily absorbed should be used.[2]][As cobalt chloride appears violet in colour, this is the colour most strongly reflected, and therefore the complementary colour of violet, which is yellow, will be most strongly absorbed, and therefore is the ideal colour for colorimetry.[3]]

- ✓ ✗ I have identified that Ryan is incorrect.[1]
- ✓ ✗ I have described the relationship between the colour used for analysis and absorbance in colorimetry.[2]
- ✓ ✗ I have described the absorption of the complementary colour of violet.[3]

d i 1.75 M

ii [This calibration curve cannot be used for the chosen solution as the absorbance is greater than the absorbance of any of the standard solutions used to produce the calibration curve.[1]] [We cannot determine the shape of the graph beyond the maximum measured value, and therefore the graph will not accurately provide information of solutions with higher concentrations than 2.0 M.[2]]

- ✓ ✗ I have compared the absorbance relative to the calibration curve.[1]
- ✓ ✗ I have explained the limitations of the graph.[2]

16 a [A monochromator is used to select a specific wavelength of light for UV-visible spectroscopy.[1]][In colorimetry, a light filter is used instead of a monochromator.[2]]

- ✓ ✗ I have identified the function of a monochromator in UV-visible spectroscopy.[1]
- ✓ ✗ I have identified the alternative to a monochromator used in colorimetry.[2]

b i [As a UV-visible spectrophotometer scans different wavelengths of light, it can be used to select the wavelength with the highest absorbance for a particular solution.[1]][Choosing the wavelength which is absorbed by the specific solution to the greatest extent enhances the accuracy of the absorbance data from the spectroscopic analysis.[2]]

- ✓ ✗ I have identified the function of scanning different light wavelengths in UV-visible spectroscopy.[1]
- ✓ ✗ I have described the advantage of the ability to choose the wavelength with the greatest absorbance.[2]

ii Blue-green light would be best suited for analysis of this solution.

c [In a solution where the metal ions do not produce a colour, a metal complex must be formed to perform UV-visible spectroscopy.[1]] [Producing a metal complex will render the colourless solution coloured such that sufficient absorbance of light for analysis can occur.[2]]

✓ ✗ I have identified that a metal complex is necessary.[1]

✓ ✗ I have explained why a metal complex is necessary.[2]

d Answers may include at least one of the following:

- Assessment of contents of blood, including bilirubin, cholesterol, iron, haemoglobin and sugars.
- Assessment of contents of urine, such as lead.
- Assessment of milk quality.
- Air quality assessment for presence of chromium and lead.
- Assessment of phosphate concentration in waterways.

17 a [An advantage of colorimetry is that it is cheap and requires little equipment.[1]][A disadvantage is that the less advanced equipment (in comparison to UV-visible spectroscopy) can lead to inaccuracies.[2]][An advantage of UV-visible spectroscopy is that it is highly accurate due to the use of a spectrophotometer including the monochromator.[3]][A disadvantage is that it is expensive due to the use of advanced equipment.[4]]

✓ ✗ I have identified an advantage of colorimetry.[1]

✓ ✗ I have identified a disadvantage of colorimetry.[2]

✓ ✗ I have identified an advantage of UV-visible spectroscopy.[3]

✓ ✗ I have identified a disadvantage of UV-visible spectroscopy.[4]

b [The correct option is III, as the amount of light absorbed by a solution is used to determine the concentration of the solution.[1]][Option I is incorrect as UV-visible spectroscopy measures the amount of light absorbed by a solution, not the amount of light emitted from it.[2]][Option II is incorrect as there are peaks not only at 210 nm, but also at approximately 400 nm.[3]]

✓ ✗ I have explained why option III is correct.[1]

✓ ✗ I have explained why option I is incorrect.[2]

✓ ✗ I have explained why option II is incorrect.[3]

c [As the solution is transparent, the complementary colour to that most strongly absorbed by the solution will be transmitted through the solution.[1]][This gives the solution the appearance of the transmitted colour.[2]]

✓ ✗ I have identified the colour that will be transmitted.[1]

✓ ✗ I have explained the appearance of the solution.[2]

Multiple lessons

18 a [Silver is considered insoluble when reacting with iodide, and therefore would form a precipitate.[1]][As a result, this decreases the amount of silver ions available in the solution to absorb light.[2]][Therefore, the concentration of silver determined by the detector would be inaccurate, and less than the true value.[3]]

✓ ✗ I have explained that silver and iodide will form a precipitate.[1]

✓ ✗ I have explained that the formation of a precipitate decreases the silver ion concentration.[2]

✓ ✗ I have connected the decreased silver concentration with the detector's readings.[3]

b $Ag^+_{(aq)} + I^-_{(aq)} \rightarrow AgI_{(s)}$

$n(Ag^+) : n(I^-) = 1 : 1$

$n(Ag^+) = 1.0 \times 0.0200 = 0.0200$ mol

$n(I^-) = 1.2 \times 0.0150 = 0.0180$ mol

$n(Ag^+) > n(I^-)$, therefore I^- is the limiting reactant

c $0.0200 - 0.0180 = 0.0020$ mol Ag^+ available to absorb light

19 a [The wavelength of light used should be approximately 600 nm.[1]][This is because only solution B absorbs 600 nm strongly, while another wavelength, such as 325 nm, would likely be absorbed by both solutions.[2]][Therefore, the absorbance of solution B could not be accurately determined at 325 nm.[3]]

✓ ✗ I have identified the appropriate wavelength for solution B.[1]

✓ ✗ I have described the absorption of solution B at different wavelengths.[2]

✓ ✗ I have identified the potential error associated with selecting an alternative wavelength.[3]

b $c(Co(NO_3)_2) = 4.1 \div 1000 = 4.1 \times 10^{-3}$ M

c i $Co(NO_3)_{2(aq)} + CuSO_{4(aq)} \rightarrow CoSO_{4(aq)} + Cu(NO_3)_{2(aq)}$

ii $n(Co(NO_3)_2) : n(CuSO_4) = 1 : 1$

$c(Co(NO_3)_2) = 4.1 \times 10^{-3}$ M $= 4.1 \times 10^{-3}$ mol per litre

$n(Co(NO_3)_2)_{10.00\ mL} = 4.1 \times 10^{-3} \times 0.010 = 4.1 \times 10^{-5}$ mol

$$c(CuSO_4) = \frac{4.1 \times 10^{-5}}{\left(\frac{20.0}{1000}\right)} = 2.05 \times 10^{-3}\ \text{M}$$

Key science skills

20 a [The non linear shape of the calibration curve is likely due to the non standardised test tubes.[1]][As some of the test tubes are transparent and some are translucent, this means that different amounts of light will be absorbed depending on which type of test tube is used.[2]]

✓ ✗ I have identified the effect of the test tubes used.[1]

✓ ✗ I have explained the different interactions between light and the test tubes.[2]

b [The use of both transparent and translucent test tubes for analysis is a systematic error.[1]][This is because the absorbance values from the standard solutions in the translucent test tubes would be consistently lower than those from the transparent tubes. In this way, there would be a constant bias in results.[2]]

✓ ✗ I have identified that this is a systematic error.[1]

✓ ✗ I have explained why this is a systematic error.[2]

c The student should ensure that all of the test tubes are transparent.

10E Analysing water using light - part 2

Theory review questions

1 A. Different amounts of energy are released as electrons return to their ground states.

2 B. The ground state is the stable energy level of an electron.

3 A. Electrons are excited when they absorb energy and enter a higher energy state.

4 A. AAS analyses the absorption of light from a sample and AES anlayses the emission of light from a sample.

5 B. Emission spectra are specific for each element and are therefore useful for qualitative analysis.

6 A. The splitting of light by a prism produces a unique spectrum for each element.

7 B. AAS calibration curves show the linear relationship between the concentration of a sample and the absorbance of light.

8 A. A monochromator selects a particular wavelength of light for analysis.

9 A. The emission spectrum is the unique spectral fingerprint of an element.

10 B. AAS involves measuring the absorption of a certain wavelength of light.

Deconstructed questions

11 C 12 B

13 [The flame test could be a replacement for atomic emission spectroscopy, however it would likely not be as effective.[1]][Firstly, atomic emission spectroscopy uses a hotter flame. This means that the probability that electrons from different elements will be excited is increased.[2]] [Secondly, atomic emission spectroscopy produces an emission spectrum specific to a particular element, enabling accurate identification of elements. There is no spectrum produced from a flame test meaning identification may be erroneous.[3]][Therefore, while the flame test and atomic emission spectroscopy are similar, these factors make atomic emission spectroscopy a more effective form of chemical analysis.[4]]

- ✓ ✗ I have compared the effectiveness of both techniques.[1]
- ✓ ✗ I have described the effect of AES's flame on electron excitation.[2]
- ✓ ✗ I have described the effect of the production of an emission spectrum on the accuracy of results.[3]
- ✓ ✗ I have identified the more effective technique.[4]

Exam-style questions

Within lesson

14 D 15 C 16 A 17 C

18 a [Like the flame test, atomic emission spectroscopy involves analysis of light emitted by electrons transitioning from a higher to a lower energy state.[1]][However, the use of a hotter flame ensures that there is sufficient energy available to excite electrons in a greater number of elements compared to the flame test.[2]][Additionally, passing the light through a prism enables an emission spectrum to be produced for each element. Since atomic emission spectroscopy produces an emission spectrum which is unique for each element, this method is significantly easier to identify a particular atom compared to a flame test, and is less subjective. For these reasons, atomic emission spectroscopy is a superior form of qualitative analysis compared to the flame test.[3]]

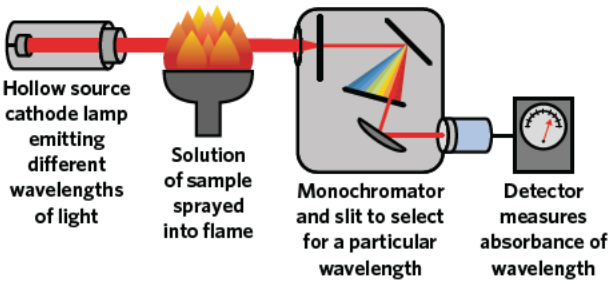
- ✓ ✗ I have identified the similarities between atomic emission spectroscopy and the flame test.[1]
- ✓ ✗ I have explained the advantage of using a hotter flame.[2]
- ✓ ✗ I have explained the advantage of passing the light through a prism and I have identified why atomic emission spectroscopy is superior to the flame test.[3]

b Emission spectra can be used for qualitative analysis, but not quantitative analysis.

c [In atomic emission spectroscopy, samples are exposed to heat energy, and light emitted from electrons as they transition from a higher energy state to the ground state provides information about the identity of the atoms in the sample.[1]][Since the emission spectra produced are used for identification purposes, this technique provides mainly qualitative information.[2]][By contrast, in atomic absorption spectroscopy a lamp emits light into a vaporised sample and, by using a monochromator to select for a wavelength of light, the amount of light at a particular wavelength that has been absorbed can be determined.[3]][The wavelength of light most readily absorbed is unique for each element, and therefore provides qualitative information about the chemical makeup of a sample. However, the absorbance value also provides information about the concentration of a given atom, thus the technique also provides quantitative information.[4]]

- ✓ ✗ I have described the process of atomic emission spectroscopy.[1]
- ✓ ✗ I have explained that AES provides mainly qualitative information.[2]
- ✓ ✗ I have described the process of atomic absorption spectroscopy.[3]
- ✓ ✗ I have explained that AAS provides qualitative and quantitative information.[4]

d



Multiple lessons

19 a Solubility refers to the extent to which a solute can be dissolved in a solvent.

b [The copper ions in copper sulphate are highly soluble in water meaning the substance is able to dissolve in water.[1]][In order for a solute to dissolve in a solvent, the intermolecular bonds formed between the solute and solvent must be sufficiently strong in order to overcome the solvent-solvent and solute-solute forces.[2]]

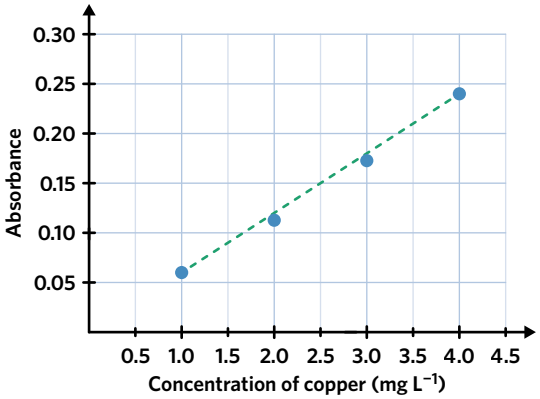
[As copper ions possess a charge, they are able to form strong intermolecular bonds with water molecules (ion-dipole bonds) enabling the dissolution of copper ions in water.[3]]

✓ ✗ I have identified that copper ions are highly soluble in water.[1]

✓ ✗ I have described the intermolecular forces involved in dissolution.[2]

✓ ✗ I have explained how the structure of copper ions facilitates their dissolution in water.[3]

c



✓ ✗ I have labelled all axes.

✓ ✗ I have constructed an accurate calibration curve.

d 2.5 mg L^{-1}

e 1 kg lines up with 1 L, therefore 1 million mg lines up with 1 L.

1 mg L^{-1} = 1 ppm

1 mg L^{-1} = 1000 ppb

1.2 mg L^{-1} = 1200 ppb

2.5 mg L^{-1} > 1.2 mg L^{-1}

Therefore there is sufficient copper sulphate in the solution for use as an anti-parasitic.

Key science skills

20 a [The absorbance value for the solution with a concentration of 0.75 M was unexpected.[1]][This is because the shape of a correct calibration curve in atomic absorption spectroscopy should be linear, however the absorbance value for this solution does not follow the linear trend.[2]]

✓ ✗ I have identified that the solution at a concentration of 0.75 M had an unexpected absorbance value.[1]

✓ ✗ I have explained why the absorbance value for this solution was unexpected.[2]

b Answers may include:

- Not all the solute was dissolved in the solvent for the 0.75 M standard solution.
- Not all the sample was vaporised by the flame.

c Answers may include:

- Same wavelength of electromagnetic radiation being assessed.
- Same equipment used for each trial.
- Samples exposed to flame for the same amount of time.
- Same temperature of the flame.

10F An introduction to acids and bases

Theory review questions

1 A. The purpose of a primary standard is to create a solution of known concentration.

2 B. The equivalence point on a titration curve is where the volume of titrant added is just enough to neutralise the analyte solution.

3 B. The end point indicates the approximate point at which the stoichiometric ratios of the reactants are met, noted by the colour change of the indicator.

4 B. Indicators make changes in the pH of a solution visible.

5 B. Burettes and conical flasks are designed to suit their functions.

6 B. Concordant titres are titres that differ by 0.10 mL from the highest value to the lowest value.

7 B The choice of indicator is based on having the end point and equivalence point of a titration as close to each other as possible.

8 B. The meniscus is the curvature of the liquid due to the difference in the forces of attraction between titrant molecules and the walls of the burete.

9 B. Strong acids and bases like NaOH and HCl are very reactive with the atmosphere and the surrounding environment.

10 A. Primary standard solutions are solutions of a precisely known concentration.

Deconstructed questions

11 A **12** D

13 Average titre = $\frac{20.00 + 20.05 + 19.95}{3}$ = 20.00 mL

$n(HCl) = c \times V = 1.0 \times 0.02000$

$= 0.020$ mol

$n(HCl) : n(NaOH)$

1 : 1

$n(NaOH) = n(HCl) = 0.020$ mol

$c(NaOH) = \frac{n}{V} = \frac{0.020}{\left(\frac{20.00}{1000}\right)} = 1.0$ M

The concentration of the sodium hydroxide present in SutherClean™ is 1.0 M

Exam-style questions

Within lesson

14 a Number of wasps = $\frac{20.00 \text{ mL}}{0.30 \text{ mL per wasp}}$ = 67 wasps

b The purpose of the indicator is to show to the experimenter the end point of the titration by changing colour.

c [These are concordant titres.[1]][This is because the difference between the highest value (22.15 mL) and lowest value (22.10 mL) is less than 0.10 mL.[2]]

✓ ✗ I have stated whether the titres are concordant or not.[1]

✓ ✗ I have justified my answer with reference to the definition of concordant titres.[2]

d Average titre = $\frac{22.10 + 22.15 + 22.10}{3}$ = 22.12 mL

$n(HCl) = 0.0010 \times \left(\frac{22.12}{1000}\right) = 0.00002212$ mol

$Mg(OH)_{2(aq)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + 2H_2O_{(l)}$

$n(Mg(OH)_2) : n(HCl) = 1 : 2$

$n(Mg(OH)_2) = \frac{1}{2} \times 0.00002122 = 0.000011$ mol

$c(Mg(OH)_2) = \frac{0.000011}{\left(\frac{20.00}{1000}\right)} = 0.00055$ M

e $n(Mg(OH)_2) = 0.00055 \times \left(\frac{0.30}{1000}\right) = 0.00000017$ mol

$n(HCl) : n(Mg(OH)_2) = 2 : 1$

$n(HCl) = 2 \times 0.00000017 = 0.0000003318$ mol

$V(HCl) = \frac{0.0000003318}{0.0010} = 0.33$ mL

15 a Traditional methods of pH measurement require aqueous solutions rather than oils.

b [The end point of the titration occurs when the indicator in the conical flask changes colour.[1]][In this case, it is when bromothymol blue transitions from a yellow to blue colour.[2]][The equivalence point is when there is the exact stoichiometric ratio of acid to base present in the conical flask.[3]][In this case, it is when the vegetable oil solution's acidity has been neutralised by the NaOH.[4]]

- I have given the definition of the end point.[1]
- I have linked the definition of the end point to the experiment.[2]
- I have given the definition of the equivalence point.[3]
- I have linked the definition of the equivalence point to the experiment.[4]

c Primary standard solutions are solutions of accurately known concentrations.[1]][They are used to standardise solutions of NaOH because NaOH is not a primary standard, and therefore, when it is made up into a solution, an accurate value for the concentration is not able to be calculated.[2]]

- I have defined a primary standard solution.[1]
- I have explained the importance of primary standard solutions in standardising a solution of NaOH.[2]

16 a [Adding more indicator would decrease the accuracy of the experiment.[1]][Since indicators are weak acids and bases, they react with the titrant and analyte in an acid-base titration.[2]][Since all the indicator has to react in order for the colour change to be observed, adding more indicator delays the onset of the end point, even when the equivalence point has been reached.[3]]

- I have identified the impact on adding more indicator to the accuracy of the experiment.[1]
- I have explained the fact that indicators react with the titrant and analyte in a reaction.[2]
- I have linked the effect of this to the equivalence point and end point.[3]

b [The best choice of indicator is when the end point, signalled by the change in colour, occurs as close as possible to the equivalence point.[1]][Since the equivalence point is 6.8, bromothymol blue will be the best choice of indicator to use.[2]][This is because its end point pH range of 6.0–7.6 and colour transition from yellow to blue perfectly matches the equivalence point of the titration reaction.[3]]

- I have defined the purpose of the indicator.[1]
- I have identified the most suitable indicator.[2]
- I have justified my answer with reference to the equivalence point and the end point.[3]

c [The five indicators mentioned above have different colours and have different pH ranges for the transition.[1]][Therefore, a rainbow of colours can be created from the mixtures of different colours and the different pH ranges that they transition in.[2]]

- I have identified the different colours of the indicators present and their different pH ranges.[1]
- I have explained how the spectrum of different colours could arise.[2]

17 a [Primary standards are required to have no water present to eliminate errors arising from weighing.[1]][Therefore, the sodium carbonate has to be dried (for example, by heating) before its use to make a primary standard solution.[2]]

- I have explained why no water can be present in primary standards.[1]
- I have identified a strategy to reduce the water content of the sodium carbonate.[2]

b Unreactive with the atmosphere, readily available and soluble in solution.

c A volumetric flask is a highly accurate piece of glassware used to create a solution of accurate volume.

d n(potassium phthalate) required for a 250.00 mL sample at 1.000 M

$= 1.0 \times \left(\frac{250.00}{1000}\right) = 0.2500$ mol

m(potassium phthalate) = 0.2500 × 204.2

= 51.05 g

1. Accurately weigh 51.05 g potassium hydrogen phthalate.
2. Transfer the potassium hydrogen phthalate into a 250.00 mL volumetric flask.
3. Add deionised water into the volumetric flask to approximately halfway.
4. Swirl the volumetric flask to dissolve the solute completely.
5. Add enough deionised water to make the solution to 250.00 mL.

Multiple lessons

18 a $CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow CH_3COONa_{(aq)} + H_2O_{(l)}$

b Carboxyl group

c Titration numbers 1, 3 and 5 produced the concordant titres.

The average titre = $\frac{21.62 + 21.58 + 21.61}{3}$ = 21.60 mL

d $n(NaOH) = 1.0 \times 0.02160 = 0.0216$ mol

$n(CH_3COOH) = n(NaOH) = 0.0216$ mol

$c(CH_3COOH) = \frac{0.0216}{\left(\frac{25.00}{1000}\right)} = 0.86$ M

e 0.86 M = mol L^{-1}

$m(CH_3COOH) = 0.86 \times (2 \times 12.0 + 4 \times 1.0 + 2 \times 16.0) = 51.6$ g/L

This means there are 51.6 grams of acetic acid per litre of vinegar in the sample.

$\frac{51.6}{1.05} = 49.14$ mL/L

$\%(v/v) = \frac{49.14}{1000} \times 100 = 4.9\%$

Therefore, the vinegar will meet the regulatory standards.

19 a A polyprotic acid is an acid that can donate multiple protons.

b [There are two equivalence points relating to both deprotonation events.[1]][Since a diprotic acid can donate a proton twice, there will be the first equivalence point when all the diprotic acid has donated the first proton, and a second equivalence point when all the deprotonated diprotic acid has donated its second proton.[2]]

✓ ✗ I have referenced the ability for diprotic acids to donate multiple protons.[1]

✓ ✗ I have explained this observation with reference to equivalence points.[2]

c [For the first equivalence point, the indicator methyl red could be used because the first equivalence point lies within methyl red's pH range (4.4–6.2).[1]][For the second equivalence point, the indicator phenolphthalein (pH range of 8.3–10.0) could be used because it transitions from colourless to pink in the pH range of the second equivalence point.[2]]

✓ ✗ I have identified a suitable indicator for the first equivalence point.[1]

✓ ✗ I have identified a suitable indicator for the second equivalence point.[2]

d n(base) = 0.01600 × 1.0 = 0.01600 mol

n(diluted acid) = n(base) = 0.01600 mol

$c(\text{diluted acid}) = \frac{0.01600}{0.02000} = 0.80$ M

Key science skills

20 a [Washing the burette with water results in some residual water being left in the burette.[1]][This means that the volume of HCl in the burette delivered will be less than what the burette says due to the water diluting the solution.[2]][Consequently, a greater volume of this diluted HCl from the burette will be required, leading to the calculated concentration of NaOH being greater than the true concentration.[3]]

✓ ✗ I have identified the impact of washing the burette with water.[1]

✓ ✗ I have described the effect of washing with water to the concentration of HCl delivered by the burette.[2]

✓ ✗ I have described the effect of washing with water on the accuracy of the calculated concentration of NaOH.[3]

b [The conical flask can have residual water because it not the concentration of analyte that is important in the flask, but the amount of analyte.[1]][Therefore, the conical flask cannot be washed with the analyte solution because residual NaOH could be left, leading to the calculated concentration of NaOH being greater than the actual concentration.[2]]

✓ ✗ I have explained why water can be used to wash the conical flask.[1]

✓ ✗ I have explained why NaOH cannot be used to wash the conical flask.[2]

c PPE should be worn like gloves and eye protection.

The lab has some easily accessible dilute base that can be used to neutralise the acid.

The broken glass is safely cleaned up.

10G Chromatography

Theory review questions

1 B. Chromatography is used both for qualitative reasons (determining and separating components in a sample) as well as quantitative reasons (determining the concentration of components in a sample).

2 B. Separation in chromatography results from the continuous movement of components between the stationary and mobile phase.

3 A. A chromatogram shows the retention times of different components in a sample as well as their absorbance peaks.

4 B. The mobile phase is the liquid which is under high pressure and passes through the chromatography column, packed with a solid stationary phase.

5 B. The retention time of a component can also depend on other factors including temperature and the length of the chromatography column.

6 B. Components will spend more time adsorbed to the phase that exhibits similar polarity.

7 A. The analysis of organic compounds (carbon-containing compounds) is one of the major uses of chromatography.

8 A. The absorbance of a component reflects the concentration at which it is present in the sample.

9 B. A calibration needs to include concentrations above and below that of the sample in question.

10 B. In chromatography, components interact with the mobile and stationary phase without being affected.

Deconstructed question

11 A **12** D

13 [Since alanine has the greatest attraction to the stationary phase, it will have the longest retention time at around 90 seconds.[1]][Glutamic acid has the greatest attraction to the mobile phase which means it will have the fastest retention time at around 55 seconds.[2]][Glycine will be the middle peak at a retention time of 70 seconds.[3]][Glutamic acid has the highest concentration in the sample because it has the largest area under its absorbance peak which is proportional to concentration.[4]]

✓ ✗ I have identified the peak responsible for alanine.[1]

✓ ✗ I have identified the peak responsible for glutamic acid.[2]

✓ ✗ I have identified the peak responsible for glycine.[3]

✓ ✗ I have linked the area under a peak to the concentration of a component in a sample.[4]

Exam-style question

Within lesson

14 C **15** A **16** D

17 a [HPLC is based upon the repeated adsorption and desorption of components between the stationary and mobile phase.[1]] [Components which are more strongly adsorbed to the stationary phase will spend more time stationary compared to those which are more strongly desorbed into the mobile phase.[2]] [Therefore, the components experience different rates of motion through the column, leading to varying retention times.[3]]

- I have identified the basic principle of HPLC.[1]
- I have explained how components have different retention times.[2]
- I have linked this idea to separation in a HPLC column.[3]

b [The scientific researcher would have to run a set of known concentrations of the specialised immune cell in the HPLC column under the same conditions.[1]] [A calibration curve can then be constructed to relate the area under the peak to the concentration of the immune cell in the patient's blood.[2]]

- I have outlined the process of creating a calibration curve.[1]
- I have explained how a calibration curve can be used to determine the concentration of the immune cell in the patient's blood.[2]

c [An increase in temperature will mean that the components in the HPLC column are less adsorbed to the stationary phase and more strongly desorbed in the mobile phase as the components are more soluble with an increased temperature.[1]] [This leads to a decreased retention time.[2]]

- I have explained the effect of temperature on the components in the sample.[1]
- I have identified the effect of temperature on retention time.[2]

d [A new calibration curve will have to be generated.[1]] [This is because a calibration curve is only accurate if it is generated with conditions identical to when the experiment with the sample of unknown concentration is carried out.[2]]

- I have stated whether a new calibration has to be generated.[1]
- I have explained the factors affecting the accuracy of a calibration curve.[2]

18 a 4

b [When a polar stationary phase is used, the component with the longest retention time will be the most polar as it is most strongly adsorbed to the stationary phase.[1]] [Therefore, P will be the most polar component.[2]]

- I have linked the length of retention time to the level of polarity of a component.[1]
- I have identified the most polar component.[2]

c [The retention times of S, P, Q, R could be matched with known compounds with the same retention times when the HPLC is carried out under identical conditions.[1]] [This is due to the unique nature of a component's retention time.[2]]

- I have explained the basis of identifying components.[1]
- I have justified why chromatography allows this due to the uniqueness of retention times.[2]

d [The coroner would have to run a set of known concentrations of ethanol in the HPLC column under the same conditions.[1]] [A calibration curve can then be constructed to relate the area under the peak to the concentration of alcohol in the patient's blood.[2]]

- I have outlined the process of creating a calibration curve.[1]
- I have explained the use of a calibration curve to determine the concentration of alcohol.[2]

Multiple lessons

19 a [Pentan-1-ol has a five carbon chain whereas ethanol has just two carbon atoms.[1]] [Since they both have the polar hydroxyl functional group, pentan-1-ol is less polar overall as the effect of the polar hydroxyl group is decreased due to a longer non-polar carbon chain. Therefore, pentan-1-ol is less polar than ethanol.[2]]

- I have described the size difference of pentan-1-ol and ethanol.[1]
- I have described the overall difference in polarity between pentan-1-ol and ethanol.[2]

b [A non-polar stationary phase has been used.[1]] [This is because the larger molecule, pentan-1-ol, which is less polar overall, has a greater retention time than the smaller, more polar molecule like ethanol.[2]] [This means that pentan-1-ol experiences a greater force of attraction to the stationary phase which means the stationary phase must be non-polar.[3]]

- I have identified the nature of the stationary phase.[1]
- I have commented on the retention times of various components.[2]
- I have linked a component's retention time and the polarity to the nature of the stationary phase.[3]

c [0.05% is a figure relating to the concentration of alcohol in the blood.[1]] [This chromatogram only gives information relating to retention time and absorbance,[2]] [and is therefore not able to reveal the concentration of ethanol in the blood as this requires a calibration curve.[3]]

- I have linked 0.05% to the concentration of alcohol in the blood.[1]
- I have described the information presented by chromatograms.[2]
- I have identified the necessity of a calibration curve when determining concentrations from chromatograms.[3]

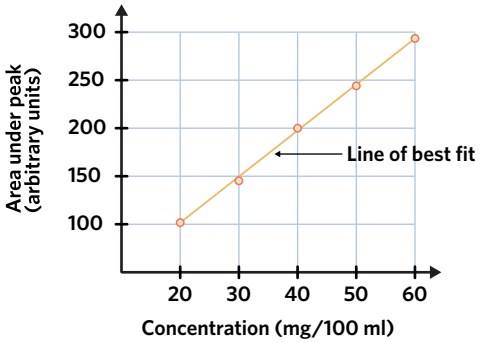
d
- It is fragile because it requires a glass column and sensitive instrumentation which might break as the booze bus moves around.
- Since conditions around Australia vary with respect to temperature, results might differ leading to inaccurate results.
- The operation of a HPLC column and interpretation of its data requires specialised chemical knowledge.

Key science skills

20 a IV: Concentration of caffeine in known solution

DV: Area under peak

b



c [The table of results show that the caffeine in Charles tea has an area under the peak of 180.[1]][Using the calibration curve, it can be seen that this correlates to a concentration of around 38 mg/100 mL.[2]] [As a result, since accuracy is a measure of how close a value is to the true value, the claim he made of 80 mg/100 ml was inaccurate.[3]]

✓ ✗	I have referred to the results from Charles' analysis.[1]
✓ ✗	I have used the calibration curve to find the caffeine concentration of Charles' tea.[2]
✓ ✗	I have made reference to the accuracy of Charles' claim.[3]

Chapter 10 review

Multiple choice questions

1 D. The majority of water on Earth is salty seawater.

2 A. As higher predators eat organisms lower in the food chain, mercury accumulates in their bodies.

3 B. Electrical conductivity testing is effective in testing salinity since conductivity is correlated with the quantity of ions in solution.

4 A. The mass of precipitate formed depends on its relationship with the limiting reagent.

5 A. A UV-visible spectrophotometer can scan the absorbance of different wavelengths of light to determine that which is most effectively absorbed.

6 D. Since the ions are not removed from the solution, the colour of the solution remains the same.

7 A. A cathode lamp composed of the metal being tested is used by AAS.

8 B. The sample of unknown concentration, the aliquot, has a set volume and is delivered via the pipette, and reacted with the titre of known concentration which is dispensed from the burette.

9 D. A high molecular weight makes the calculation of the number of moles more accurate.

10 A. Using the calibration curve, a y-value of 6.0 corresponds with an x-value of approximately 1.05.

Short answer questions

11 a

Term	Definition
Equivalence point	The point at which the reactants have reacted completely in accordance with the stoichiometric ratio of the balanced equation.
End point	The point at which the indicator undergoes a permanent colour change.
Concordant titres	Titres within a range of 0.10 mL.

b i $C_2H_2O_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2C_2O_{4(aq)} + 2H_2O_{(l)}$

ii $n(NaOH) = 1.0 \times 0.025 = 0.025$ mol

$n(NaOH) : n(C_2H_2O_4) = 2 : 1$

$n(C_2H_2O_4) = \frac{1}{2} \times 0.025 = 0.0125$ mol

$c(C_2H_2O_4) = \frac{0.0125}{0.0274} = 0.47$ M

12 a i [The ions under analysis should be the limiting reactant.[1]] [This is because, it is essential that all the ions react with the precipitating agent so its quantity can be accurately measured.[2]]

✓ ✗	I have identified that the ions in the water sample are the limiting reactant.[1]
✓ ✗	I have explained that all of the ions in the water sample must be measured.[2]

ii $n(Pb(NO_3)_2) : n(CaCl_2) = 1 : 1$

$n(Pb(NO_3)_2) = \frac{6.34}{207.2 + 2 \times (14.0 + 3 \times 16.0)} = 0.0191$ mol

$n(CaCl_2) = \frac{2.51}{40.1 + (2 \times 35.5)} = 0.0226$ mol

$n(Pb(NO_3)_2) < n(CaCl_2)$

$Pb(NO_3)_2$ is the limiting reactant.

iii $n(PbCl_2) = 0.0191$ mol

$m(PbCl_2) = 0.0191 \times [207.2 + (2 \times 35.5)] = 5.33$ g $PbCl_2$ formed.

b [The electrical conductivity in the river water sample is greater since it has a higher concentration of ions than pure water.[1]][This is because water does not conduct electricity well, whereas ions are effective conductors of electricity.[2]][The electrical conductivity of a solution increases as the concentration of ions increases.[3]]

✓ ✗	I have identified the main reason for the difference in electrical conductivity.[1]
✓ ✗	I have explained the difference in conductivity between water and ions.[2]
✓ ✗	I have explained the effect of increasing ion concentration on electrical conductivity.[3]

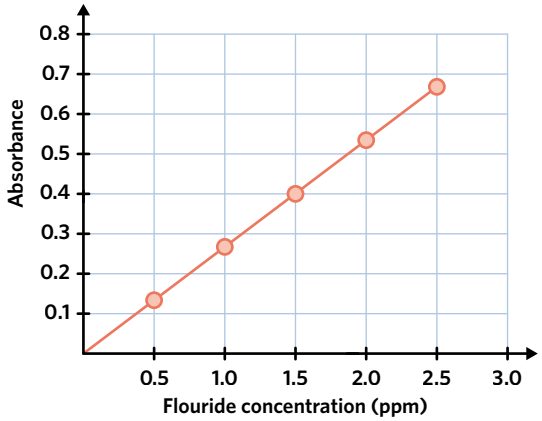
13 a Potential sources of contamination include:

- Industrial run-off
- Agricultural run-off
- Mining operations
- Sewage

b [Colorimetry involves passing light of a particular wavelength through a sample and using a detector to determine the amount of light absorbed. The absorbance data is used to determine the concentration of the substance.[1]][On the other hand, atomic absorption spectroscopy involves passing of light from a cathode lamp through a vapourised sample. A wavelength of light is then selected using a monochromator and the absorbance by the sample is measured and concentration established.[2]][Therefore, both colorimetry and AAS utilise the absorbance of light to obtain qualitative and quantitative data about a sample.[3]][However, in colorimetry the wavelength of light is chosen before passing it through the sample. Whereas, in AAS the monochromator selects the wavelength after the sample has been exposed to light. Furthermore, the sample is vapourised in AAS but not in colorimetry.[4]]

- ✓ ✗ I have described the method of colorimetry.[1]
- ✓ ✗ I have described the method of atomic absorption spectroscopy.[2]
- ✓ ✗ I have identified the similarities between colorimetry and AAS.[3]
- ✓ ✗ I have identified the differences between colorimetry and AAS.[4]

c i



ii [For the Yenisei River, the absorbance value of 0.37 corresponds to a fluoride concentration of 1.4ppm. Since this is less than the 1.5 ppm limit, the water in this river is safe to consume.[1]]
[For the Yangtze River, the absorbance value of 0.64 corresponds to a fluoride concentration of 2.4 ppm. Since this is greater than the 1.5 ppm limit, it is unsafe to drink the water.[2]]

- ✓ ✗ I have explained that the Yenisei River water is safe to consume.[1]
- ✓ ✗ I have explained that the Yangtze River water is unsafe to consume.[2]

14 a The presence of four peaks on the chromatogram means there are four substances in the sample.

b The retention time is the amount of time taken for a component to travel through a chromatography column.

c [In this HPLC, the stationary phase used is non-polar meaning that non-polar substances would have longer retention times than polar compounds.[1]][Since W had the shortest retention time, it would be the most polar component present.[2]]

- ✓ ✗ I have explained the relationship between the non-polar stationary phase and retention time.[1]
- ✓ ✗ I have identified the most polar component as the one with the lowest retention time.[2]

d [Component W has the highest concentration[1]][as it has the greatest relative peak area.[2]]

- ✓ ✗ I have identified the component with the highest concentration.[1]
- ✓ ✗ I have identified the relationship between peak area and concentration.[2]

e

Change to conditions	Effect on retention time
Longer column	Increased
Increased flow rate of mobile phase	Decreased
Decreased temperature	Increased

Key science skills questions

15 a Quantitative data.

b 23.70 − 2.50 = 21.20 mL
28.60 − 7.35 = 21.25 mL
22.25 − 1.60 = 20.65 mL
[The titres have a range of 0.65 mL.[1]][Concordant titres have a range of 0.10mL. Since 0.65 mL > 0.10 mL, the titres are not concordant.[2]]

- ✓ ✗ I have included calculations in my answer.
- ✓ ✗ I have identified the range of the titres recorded.[1]
- ✓ ✗ I have identified the concordant nature of the titres.[2]

c [An example of a systematic error is consistently washing the burette or pipette with the wrong solution after use.[1]][An example of a random error is reading the burette value incorrectly.[2]]

- ✓ ✗ I have identified a systematic error that may have produced non-concordant titres.[1]
- ✓ ✗ I have identified a random error that may have produced non-concordant titres.[2]

GLOSSARY

A

Accuracy how close measured values are to their true value p. 26

Acid species that donates a proton, H^+, during an acid-base reaction p. 392

Acid-base reaction chemical reaction where a proton, H^+, is transferred between two chemical species p. 392

Acidic solution hydronium is more concentrated than hydroxide: $[H_3O^+] > [OH^-]$ p. 407

Activity rate at which an enzyme converts between reactants and products p. 430

Addition polymerisation reaction that joins monomers together to form a polymer p. 299

Addition reaction one molecule combining with another molecule to form a larger molecule p. 246

Adenosine triphosphate (ATP^{3-}) high-energy organic molecule used to drive cellular reactions p. 481

Adsorption attachment of a molecule to a surface p. 52

Adsorption (chromatography) component adheres to the stationary phase p. 563

Aim purpose of an experiment p. 2

Alcohol organic compound that has a hydroxyl (-OH) group bonded to a carbon atom in the carbon chain p. 281

Aliquot an exact amount of a sample solution that is to be analysed during titration p. 546

Alkali base that is soluble in water p. 392

Alkane hydrocarbon that contains only single carbon to carbon bonds and without functional groups p. 246

Alkene hydrocarbon with at least one double carbon to carbon bond p. 246

Alkyl group branch associated to the parent carbon chain that contains only carbon and hydrogen atoms p. 259

Alkyne hydrocarbon with at least one triple carbon to carbon bond p. 246

Allotrope one of the different physical forms an element can exist in due to different structural arrangements of atoms p. 234

Alloy material that is made by combining a metal with another chemical element which can be a metal or a nonmetal p. 142

Alloying agent element combined with a main metal to make an alloy p. 142

Amino acid organic molecule with a central carbon attached to an amino group, a carboxyl group, a side chain (R group) and a hydrogen atom p. 430

Amino group functional group with molecular formula NH_2 p. 430

Amorphous region region of polymer with random entanglement of chains p. 307

Amount number of moles of substance p. 98

Amphiprotic species that can either donate or accept a proton, H^+ p. 392

Analyte substance which is quantified or identified p. 516

Anion negative ion formed due to the addition of electron(s) to an atom p. 164

Annealing heat treatment in which metals are heated and left to cool down slowly p. 142

Aqueous solution solution in which the solvent is water p. 324

Atactic side chains are randomly distributed along the polymer p. 307

Atom smallest unit of matter p. 52

Atomic absorption spectroscopy (AAS) method of chemical analysis using the light absorbed by metals to determine the type and number of metals present in a sample p. 538

Atomic emission spectroscopy (AES) method of chemical analysis using the intensity of light emitted from a flame to determine the metals present p. 538

Atomic mass unit (amu) unit of mass used to express atomic masses, relative to $\frac{1}{12}$ of the mass of a single carbon-12 atom p. 88

Atomic number number of protons in the nucleus of an atom p. 66

Atomic radius size of an atom measured as the distance from the nucleus to the valence shell in picometers (pm) p. 72

ATP synthase enzyme complex that uses protons stored in the mitochondrion's intermembrane space to drive the synthesis of ATP^{3-} p. 481

Aufbau principle rule that states subshells are filled by electrons from the lowest to the highest energy level p. 52

Autoionisation process by which an atom or molecule spontaneously forms ions p. 407

Average titre mean value of the concordant titres p. 546

Avogadro's number (N_A) number of atoms in 12 g of ^{12}C, equal to 6.02×10^{23} p. 98

B

Balanced equation same number of atoms of each element on either side of the chemical equation p. 113

Ball-and-stick model representation of a molecule demonstrating the three-dimensional arrangement of atoms p. 196

Base species that accepts a proton, H^+, during an acid-base reaction p. 392

Basic solution hydroxide is more concentrated than hydronium: $[H_3O^+] < [OH^-]$ p. 407

Bias difference between the average of a large set of measurements and the true value p. 26

Bioaccumulation gradual accumulation of substances within an organism p. 505

Boiling point temperature at which a liquid transitions into a gas p. 216

Brittle tendency of a material to fracture (break) when subjected to pressure p. 171

Brønsted-Lowry theory definition of an acid as a chemical species which donates a proton, H^+, to a base p. 392

Burette glassware which delivers accurate, yet variable volumes of titrant p. 546

C

Calibration curve graph depicting the relationship between concentration and light absorbance for a particular substance p. 528

Carbon chain linear set of carbon atoms in a hydrocarbon p. 246

Carbonic anhydrase enzyme that increases the rate of the reaction $CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$ p. 430

Carboxyl functional group functional group that has a carbon double bonded to an oxygen and also singly bonded to a hydroxyl group p. 281

Carboxylic acid organic compound that contains a carboxyl functional group (-COOH) p. 281

Catalyst substance that increases the reaction rate without itself being consumed or permanently changed p. 291

Cation positive ions formed when an atom loses its valence electron(s) p. 132

Cellular respiration process by which cells create energy from biochemical reactions p. 481

Chemical symbol abbreviation used to represent a chemical element p. 66

Chromatogram visual output of chromatography which can be analysed p. 563

Chromatography analytical technique used to determine the identity of components in a mixture as well as their concentration p. 563

Coenzyme Q mobile electron carrier in the ETC that receives electrons from NADH and $FADH_2$ p. 481

Colorimeter instrument used to undertake colorimetric analyses p. 528

Colorimetry method of chemical analysis exploiting the light absorbance properties of substances with different concentrations p. 528

Complementary colours colours opposite to each other on the colour wheel p. 528

Compound two or more atoms of different elements bonded together p. 66

Concentrated solution solution with relatively high amount of solute compared to the volume of solution p. 352

Concentration amount of chemical substance in a unit of volume p. 546

Concordant titres volume of three or more titres that fall within approximately 0.1 mL of each other p. 546

Condensation reaction chemical reaction where two molecules combine with the release of a water molecule p. 291

Conductive polymers polymers able to conduct electricity p. 307

Conical flask glassware which the analyte is placed in p. 546

Conjugate acid chemical species formed from a base that has accepted a proton, H^+ p. 392

Conjugate acid-base pair acid and its conjugate base or a base and its conjugate acid p. 392

Conjugate base chemical species formed from an acid that has donated a proton, H^+ p. 392

Conjugate redox pair electron acceptor (oxidising agent) and its corresponding electron donor p. 448

Contaminant substance that makes another substance impure and is usually toxic p. 505

Control group (in chemistry) sample that is subjected to the same conditions as all other samples without the independent variable p. 18

Controlled variable(s) variable(s) held constant throughout the experiment p. 2

Copolymer polymers composed of two or more types of monomers p. 307

Core charge attractive force between the positive nucleus and valence electrons, measured as the difference in the number of protons and inner electrons p. 72

Covalent layer lattice atoms covalently bonded together in a two-dimensional arrangement to form layers, held together by weak dispersion forces p. 234

Covalent network lattice regular three-dimensional arrangement of atoms covalently bonded together extending outwards p. 234

Crude oil unrefined petroleum deposit composed of hydrocarbons p. 246

Crystalline region region of polymer with chains tightly packed together p. 307

Cytochrome c mobile electron carrier in the ETC with a central Fe^{3+} ion capable of accepting one electron p. 481

Cytoplasm thick, jelly-like liquid that fills up the space within a cell p. 481

D

Delocalised electrons electrons that freely move between metal cations in the metallic bonding model p. 132

Dependent variable variable that is measured by the experimenter p. 2

Deprotonation loss of a proton, H^+, from an acid p. 392

Desorption a component is released from the stationary phase and dissolves into the mobile phase (in the context of chromatography) p. 563

Diatomic molecule molecule consisting of only two atoms (they may be different elements) p. 196

Dichromate ion polyatomic ion with molecular formula $Cr_2O_7^{2-}$ p. 458

Dilute (adj.) solution with decreased concentration of solute p. 352

Diluted solution solution with relatively low amount of solute compared to the volume of solution p. 352

Dilution process of decreasing concentration of a solution by adding more solvent, usually water p. 546

Dilution factor ratio of volume of stock solution to the volume of final (diluted) solution p. 546

Dipole-dipole force force of attraction between a partial positive charge on one molecule and a partial negative charge on another molecule p. 216

Diprotic species that donates two protons, H^+, during ionisation p. 392

Dispersion force electrostatic force of attraction between molecules resulting from instantaneous dipole moments p. 216

Dissociate (ions) process where ions break away from their lattice p. 344

Dissociation process by which a molecule separates into smaller particles p. 392

Dissolution substance dissolves in another substance p. 344

Dissolve substance becomes incorporated into a liquid to form a solution p. 344

Double covalent bond bond formed by the sharing of four electrons/two pairs of electrons (two covalent bonds) p. 196

Double displacement reaction reaction between ionic compounds that results in the exchange of ions to produce new compounds p. 178

Ductility ability to be hammered or stretched into a thin shape without breaking p. 132

E

Economic considerations taking into account the associated costs (e.g. time and money) p. 9

Elastomer polymers that form occasional cross-links and are elastic p. 299

Electrical conductivity ability to allow an electric current to flow through p. 132

Electromagnetic radiation waves of the electromagnetic field including visible and ultraviolet light p. 528

Electron light negatively charged particle that exists outside the nucleus of an atom p. 52

Electron configuration arrangement of electrons in shells and/or subshells p. 196

Electron dot formula representation of a molecule where valence electrons are shown as dots around the chemical symbol p. 196

Electron transport chain (ETC) series of protein complexes on the inner mitochondrial membrane that accept electrons from NADH and $FADH_2$ to reduce oxygen to water p. 481

Electronegativity how strongly an atom attracts electrons towards itself p. 206

Electrostatic force of attraction attractive force between charged particles p. 132

Element pure substance containing only a single type of atom p. 66

Eluent combination of the mobile phase and sample that moves through the HPLC column p. 563

Elute (used for HPLC) movement of a substance out of the chromatography column p. 563

Emission spectra band of frequencies emitted by an atom due to electrons returning to the ground state from an excited state p. 52

Emission spectrum spectrum of electromagnetic wavelengths emitted by a metal ion in atomic emission spectroscopy p. 538

Empirical formula simplest, whole number ratio of atoms of different elements in a compound p. 113

End point point at which a colour change occurs during titration p. 546

Energy shells or energy level orbits, containing different levels of energy, around the nucleus of an atom where electrons are found according to the Bohr model p. 52

Environmental considerations taking into account the environmental implications p. 9

Enzyme proteins capable of increasing the rate of a chemical reaction in a biological system p. 430

Equivalence point point at which the reaction has occurred as per the stoichiometric ratio of the balanced equation p. 546

Error difference between the measured value and the true value p. 26

Ester organic compound that contains an ester functional group (-COO-) p. 291

Esterification reaction chemical reaction between organic compounds that forms at least one ester as a product p. 291

Ethical considerations taking into account the effect on other living organisms (e.g. humans and animals) p. 9

Eutrophication richness of nutrients in water system leading to the growth of plant-life p. 505

Excited electron electron which has absorbed energy and is elevated to a higher unstable energy level p. 538

Excited state temporary energy level that an electron occupies once it has absorbed energy p. 52

F

First ionisation energy energy required to remove the first valence electron from an atom p. 72

Flavin adenine dinucleotide (FAD) electron carrier used in cellular respiration with reduced form $FADH_2$ p. 481

Fossil fuel fuels derived from fossilized plant and animal matter p. 246

Fractional distillation separating of a liquid mixture into fractions differing in boiling point p. 246

Full equation chemical equation with all aqueous ions and chemical species present p. 376

Functional group specific groups of atoms within a compound that affect the properties of the compound p. 281

G

Glucose simple sugar molecule with molecular formula $C_6H_{12}O_6$ p. 481

Gravimetric analysis chemical analysis involving the weighing of a precipitate formed from an analyte p. 516

Ground state lowest energy level that an electron occupies p. 52

Group 16 hydrides Group 16 element covalently bonded to hydrogen p. 324

Groups columns in the periodic table p. 72

H

Half-reaction reduction or oxidation equation of a complete redox reaction p. 448

Haloalkane alkane possessing a halogen functional group; a halogenated alkane p. 246

Halogen element in group 17 of the periodic table p. 246

Halogenation replacement of a hydrogen atom with a halogen in a molecule p. 246

Hardness resistance to deformation when subjected to pressure p. 171

Hardness of water amount of calcium and magnesium dissolved in water p. 505

Heat capacity relationship between the heat absorbed by a substance and its temperature change p. 335

Heat conductivity ability to allow heat to pass through p. 132

Heavy metals metals with relatively high atomic masses p. 505

High density polyethene tightly packed polymer produced at low pressures p. 299

Homogenous same proportion of particles throughout the solution (i.e. well-mixed) p. 324

Homologous series series of hydrocarbons that have a similar chemical structure and chemical properties p. 246

HPLC (High performance liquid chromatography) a type of chromatography technique which pumps the mobile phase through a very tightly packed stationary phase under high pressure p. 563

Hydrated (ions) ions that become surrounded by water molecules p. 324

Hydrocarbon compounds consisting of hydrogen and carbon p. 246

Hydrogen pop test test used to indicate a reaction between a metal and an acid p. 150

Hydrolysis reaction chemical reaction where water is used to break the bonds of a substance p. 291

Hydronium ion positively charged chemical species with molecular formula H_3O^+ p. 392

Hydroxide ion negatively charged chemical species with molecular formula OH^- p. 392

Hydroxyl group functional group consisting of an oxygen atom covalently bonded to a hydrogen atom (-OH) p. 281

Hypothesis testable statement which predicts the outcome of an experiment p. 2

I

Independent variable variable that is deliberately manipulated by the experimenter p. 2

Indicator substance that will experience a colour change under a specific condition p. 407

Inner shell electrons electrons not in the valence shell p. 196

Insoluble unable to be dissolved in a given solvent p. 171

Instantaneous dipole moment partial positive and negative charge that suddenly appears in an atom/molecule due to the random movement of electrons p. 216

Interstitial alloy alloy in which alloying agent atoms stay in the gaps between main metal cations p. 142

Intramolecular bonds bonding between atoms within a molecule p. 196

Ion charged atom p. 72, 164

Ion-dipole attraction electrostatic attraction between a molecule's permanent dipole and an ion p. 324

Ionic bond electrostatic force of attraction between cations and anions p. 164

Ionic compound compound made up of ions held together by electrostatic forces of attraction p. 164

Ionic equation equation of a chemical reaction that does not include spectator ions p. 420

Ionic product expression of the ions present in water at 25°C given by $K_W = [OH^-][H_3O^+] = 10^{-14}\ M^2$ p. 407

Ionic salt ionic compound formed from an anion and cation p. 420

Ionisation process by which a chemical species gains or loses an electric charge p. 392

Isomers compounds that share the same molecular formula but different structure p. 246

Isotactic all side chains are on one side of the polymer p. 307

Isotope variants of an element which have the same atomic number but a different number of neutrons in their nuclei p. 66

K

Kinetic energy energy of an object due to its motion p. 216

L

Latent heat of fusion amount of energy required to convert 1 mol of a substance from its solid state to its liquid state at the melting point of the substance p. 335

Latent heat of vaporisation amount of energy required to convert 1 mol of a substance from its liquid state to its gaseous state at the boiling point of the substance p. 335

Lattice regular three-dimensional arrangement of atoms p. 52

Law of conservation of mass law that states that in a closed system, the total mass of reactants equals the total mass of products p. 516

Lone pair of electrons pairs of valence electrons not shared with another atom in a covalent bond p. 196

Low density polyethene branched polymer produced at high pressures p. 299

Lustre shiny and glossy appearance p. 132

M

Malleability ability to deform under pressure without breaking p. 132

Mass number mass of an atom, approximately equal to the average sum of protons and neutrons p. 66

Mass spectrometry analytical technique used to measure the mass of ions relative to their charge p. 88

Mass-mass stoichiometry calculation of reactants and products in a chemical reaction using the law of conservation of mass p. 516

Mass-to-charge ratio the mass of an ion divided by its charge p. 88

Material Safety Data Sheet (MSDS) document that outlines the health and safety information associated with different materials and chemicals p. 9

Melting point temperature at which a substance changes its state from solid to liquid p. 171

Meniscus curved upper surface of liquid p. 546

Metabolism process by which the body breaks down large molecules and harnesses their energy p. 481

Metal carbonate compound of a metal cation and the carbonate ion, CO_3^{2-} p. 420

Metal displacement reaction redox reaction in which a more reactive metal displaces a less reactive metal's cation from solution p. 472

Metal hydrogen carbonate compound of a metal cation and a bicarbonate ion, HCO_3^- p. 420

Metal hydroxide compound of a metal cation and the hydroxide ion, OH^- p. 420

Metal oxide formed as a result of a reaction between a metallic element and oxygen p. 150

Metallic bonding electrostatic force of attraction between delocalised electrons and cations in a metallic lattice structure p. 132

Metallic character chemical properties associated with metals p. 81

Metalloids elements which possess properties that are in between those of metals and nonmetals p. 81

Mitochondria enclosed organelles that are the site of energy production within a cell p. 481

Mobile phase phase that flows through the stationary phase carrying components of a sample with it p. 563

Molar mass mass, in grams, of one mole of any substance p. 104

Molarity (M) concentration of a solution measuring the number of moles of solute per litre of solution p. 352

Mole (*n*) unit of measurement defined as 6.02×10^{23} chemical entities p. 98

Molecular formula chemical formula which gives the type and number of atoms of each element in a molecule p. 88

Molecular substance substance with two or more atoms joined by covalent bonds p. 216

Molecule two or more atoms chemically bonded together p. 196

Monatomic ions ions made up of only one atom p. 178

Monochromator device used to select a particular wavelength of light p. 528

Monomer molecules that can react with other molecules to form larger molecules (polymers) p. 299

Monoprotic species that donates one proton, H^+, during ionisation p. 392

N

Nanoparticles particles on the 10^{-9} m scale p. 52

Nanoscale measurements on the 10^{-9} m scale p. 52

Net ionic equation full equation once the spectator ions have been removed showing only the chemical species that have taken part in a reaction p. 376

Network atoms bonded in a continuing arrangement extending outwards p. 234

Neutral solution hydronium and hydroxide concentrations are equal: $[H_3O^+] = [OH^-]$ p. 407

Neutralisation reaction reaction between an acid and a metal hydroxide to produce water p. 420

Neutron dense neutrally charged particle inside the nucleus of an atom p. 52

Nicotinamide adenine dinucleotide (NAD^+) electron carrier used in cellular respiration with reduced form NADH p. 481

Noble gases unreactive gases found in group 18 of the periodic table p. 81

Non-polar covalent bond covalent bond where the electrons are shared equally p. 206

Non-renewable fuel fuel consumed at a rate greater than it is replenished p. 246

Non-scientific ideas ideas that are not developed by following the scientific method p. 2

Nucleons subatomic particles found in the nucleus p. 52

Nucleus region at the centre of an atom that contains protons and neutrons p. 52

O

Octet rule general principle which states that atoms like to have 8 electrons in their valence shell for stability p. 196

Optimal pH pH at which an enzyme will be most effective p. 430

Orbital region around the nucleus of an atom where electrons are likely to be found p. 216

Orbitals regions with the highest probability of finding electrons p. 52

Ore deposit in Earth's outermost layer containing metals and other minerals p. 132

Organic chemistry chemistry of carbon-containing compounds p. 196

Organic compounds compounds consisting of carbon atoms most commonly covalently bonded to hydrogen, nitrogen, phosphorus or oxygen atoms p. 246

Organometallic compounds organic compounds containing at least one metallic element in its structure, with a carbon-metal bond p. 505

Oxidation chemical reaction where a chemical species loses electrons p. 448

Oxidation number number assigned to an atom in a substance used to determine the movement of electrons in redox reactions p. 458

Oxidised word that describes a chemical species that has undergone oxidation p. 448

Oxidising agent (oxidant) chemical species that causes another substance to undergo oxidation and is itself reduced p. 448

P

Parent chain longest continuous chain of carbon-carbon bonds. This needs to include any carbon double/triple bonds (ie functional groups) if present p. 259

Parent name component of name that refers to the length of the parent chain p. 259

Parts per billion (ppb) one part in one billion parts of a solution p. 352

Parts per million (ppm) one part in one million parts of a solution p. 352

Pauli exclusion principle rule that states an orbital can't hold more than two electrons p. 52

Peptide/amide bond functional group that links amino acids together with molecular formula CONH p. 430

Percentage composition percentage by mass of an element in a compound p. 113

Periodic table table of chemical elements in which elements are arranged in order of increasing atomic number p. 66

Periodicity characteristics of elements in a period p. 72

Periods rows in the periodic table p. 72

Permanent dipole moment covalent molecules with a permanent positive and negative charge due to the difference in electronegativity between the atoms p. 206

Permanganate ion polyatomic ion with molecular formula MnO_4^- p. 458

Personal error mistake or misinterpretation of a reading or method by the experimenter p. 26

pH measure of the concentration of H_3O^+ ions in a solution using the formula $pH = -\log_{10}[H_3O^+]$ p. 407

pH scale values of pH arranged on a logarithmic scale p. 407

Pipette laboratory tool to measure and deliver small volumes of solutions accurately p. 546

Polar covalent bond covalent bond between atoms with an unequal distribution of electrons in the bond p. 206

Polar molecule molecule with a partial positively and negatively charged end p. 206

Polarity property of having a partial positive and negative charge p. 206

Polyatomic ion ions made up of a group of atoms covalently bonded together p. 178

Polymer molecule produced from the addition reaction between monomers p. 299

Polyprotic species that donates multiple protons, H^+, during ionisation p. 392

Potable alternative term for drinkable p. 500

Precipitate solid compound formed in a precipitation reaction p. 376

Precipitation reaction chemical reaction which results in the formation of a solid from an aqueous solution p. 376

Precision how close measured values are to each other p. 26

Primary data original data collected firsthand by researchers p. 9

Primary source source that provides the original raw data p. 2

Primary standard substance used to make a primary standard solution p. 546

Primary standard solution solution whose concentration can be precisely calculated p. 546

Protein complex organic molecules formed from multiple amino acids p. 430

Proton light positively charged particle inside the nucleus of an atom p. 52

Q

Qualitative data non-numerical (descriptive) data collected based on observations taken during an experiment p. 9

Quantitative analysis technique that identifies the amount of substance present p. 9

Quantitative data numerical data collected during experiments p. 9

Quantum mechanics field of study which looks at the behaviour of very small objects p. 52

Quenching heat treatment in which metals are heated and left to cool down quickly p. 142

R

Random error error in measurement that differs in amount each time the experiment is conducted and is usually a 'one-off' error p. 26

Ratio numerical relationship between two or more elements p. 113

Reactant (in the context of precipitation) substance that reacts with an analyte to form a precipitate p. 516

Reactive metal metal that will readily react with the H^+ ions from an acid p. 420

Reactivity tendency of atoms to undergo a chemical reaction p. 81

Reactivity series of metals organised depiction of metals and their cations, ranked according to their strength as reducing and oxidising agents p. 472

Redox reaction chemical reaction that involves the transfer of electrons between two chemical species p. 448

Reduced word that describes a chemical species that has undergone reduction p. 448

Reducing agent (reductant) chemical species that causes another substance to undergo reduction and is itself oxidised p. 448

Reduction chemical reaction where a chemical species gains electrons p. 448

Relative abundance percentage of a particular isotope found in a naturally occurring sample of an element p. 88

Relative atomic mass (A_r) weighted mean of the relative isotopic masses of all isotopes of an element p. 88

Relative formula mass mass of a unit of a substance, particularly an empirical unit, relative to $\frac{1}{12}$ of the mass of a single carbon-12 atom p. 88

Relative isotopic mass mass of a single atom of an isotope of an element relative to $\frac{1}{12}$ of the mass of a single carbon-12 atom p. 88

Relative molecular mass (M_r) average mass of one molecule of a substance relative to $\frac{1}{12}$ of the mass of a single carbon-12 atom p. 88

Renewable fuel fuel able to be replenished at a rate greater than or equal to its rate of consumption p. 246

Reservoir a water storage location p. 500

Retention time, R_t measure of time taken for a component to travel through a chromatography column eg HPLC column p. 563

Rust orange or brown metal oxide buildup on metals caused by oxygen p. 150

S

Saturated hydrocarbon hydrocarbon possessing only single bonds between carbon atoms p. 246

Scale range of values used for measuring size or other attributes p. 52

Scientific ideas ideas that are developed by following the scientific method p. 2

Scientific method procedure used to investigate scientific ideas p. 2

Secondary data data that has been previously collected that is now accessible to different researchers p. 9

Secondary source source that has interpreted primary sources p. 2

Significant figures number of digits required to express a number to a certain level of accuracy p. 26

Single covalent bond bond formed by the sharing of two electrons/one pair of electrons p. 196

Social considerations taking into consideration the effect on others p. 9

Solar energy energy produced by the sun p. 500

Solubility extent to which a substance dissolves in a solvent p. 364

Soluble able to be dissolved in a given solvent p. 171

Solute substance that dissolves in a solvent p. 324

Solution mixture where a solute has dissolved in a solvent p. 324

Solvent substance that dissolves another substance, usually water p. 324

Space-filling model representation of a molecule demonstrating the three-dimensional arrangement, and the respective shape and size of atoms p. 196

Specific heat capacity energy (J) required to raise the temperature of 1 gram of a given substance by 1°C p. 335

Spectator ion aqueous ions which do not react and remain dissolved in solution throughout a reaction p. 420

Spectator ions aqueous ions which do not react and remain dissolved in solution throughout a reaction p. 376

Spectroscopy analysis of the interaction between matter and electromagnetic radiation p. 528

Spontaneous redox reaction redox reaction that occurs without the addition of any extra heat or energy p. 472

Standard solution solution of accurately known concentration p. 528

Stationary phase solid onto which the components of a sample adsorb p. 563

Stoichiometry the relationship between relative quantities of reactants and products in a chemical reaction p. 516

Strength how readily an acid or base will donate or accept a proton, H^+ p. 392

Structural isomers isomers that have the same molecular formula but the atoms are arranged in different orders p. 281

Subatomic particle particles that exist inside the atom p. 52

Substituents atom or group of atoms replacing one or more hydrogens on the parent chain of a hydrocarbon p. 246

Substitution reaction one molecule replacing a part of another molecule p. 246

Substitutional alloy alloy in which alloying agent atoms replace some main metal cations p. 142

Syndiotactic side chains on alternating sides of the polymer p. 307

Systematic error error in measurement by the same amount in the same direction every time which also includes errors that are inherent in the experiment p. 26

Systematic name (referring to organic compounds) standardised name p. 259

T

Tempering heat treatment in which metals are heated and left to cool down in the air p. 142

Tetrahedron shape with four triangular faces p. 206

Thermoplastic polymer polymer with relatively weak intermolecular bonds that can be remoulded at a particular temperature p. 299

Thermosetting polymer polymer with relatively strong intermolecular bonds that degrades when heated at a particular temperature and cannot be remoulded p. 299

Titrant solution of known concentration used in a titration reaction p. 546

Titration quantitative technique used to find the unknown concentration and/or amount of substance in a solution p. 546

Titre volume of the solution delivered from the burette to reach the end point of a titration p. 546

Triple covalent bond bond formed by the sharing of six electrons/three pairs of electrons (three covalent bonds) p. 196

Triprotic species that donates three protons, H^+, during ionisation p. 392

U

Uncertainty level of doubt regarding the data obtained (for VCE purposes, we are only looking at this from a qualitative perspective) p. 26

Unpaired electron electron in an orbital by itself p. 196

Unsaturated hydrocarbon hydrocarbon possessing at least one double or triple bond between carbon atoms p. 246

UV-visible spectrophotometer instrument used in UV-visible spectroscopy p. 528

V

Valence electrons electrons in the outermost shell of an atom p. 196

Valence shell outermost energy shell p. 72

Valence shell electron pair repulsion theory (VSEPR) model used to predict the different shapes of molecules based on the extent to which electron pairs repel each other p. 206

Valence structure representation of a molecule where chemical symbols are written (as well as valence electrons), joined by single, double or triple lines, corresponding to the type of bond p. 196

Validity whether or not the experiment and its components including the results address the aim and hypothesis of the research p. 26

Volume amount of space occupied by a substance p. 546

Volumetric analysis quantitative analytical technique used to determine unknown concentrations of solutions p. 546

W

Water cycle the continuous movement of water on Earth as it transitions between solid, liquid and gaseous states p. 500

Symbols

%(m/m) mass percent concentration of solute in solution p. 352

%(m/v) mass percent concentration of solute per unit of volume of solution p. 352

%(v/v) volume percent concentration of solute in solution p. 352

%(w/w) weight percent concentration of solute in solution p. 352

ACKNOWLEDGEMENTS

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RESEARCH QUESTIONS FOR UNIT 1 AOS 3

The information below has been taken from your study design and outline the possible options for your research investigation in Unit 1 AOS 3.

Option 1: The origin of the elements

Questions that may be explored in this investigation include:

- How are atoms 'seen'?
- Are there more elements to be discovered?
- What are electrons, protons and neutrons made of?
- How do we know what elements are in the Universe?
- What is the evidence that living things are made of stardust?
- Why are the ten most abundant elements in the Universe not the same as the ten most abundant elements on Earth?
- How does the abundance of elements on Earth compare with the abundance of elements in humans?
- Would there be life if elements did not form compounds?
- Why are the nuclei of some elements unstable?

Option 2: The development of the periodic table

Questions that may be explored in this investigation include:

- Is alchemy chemistry?
- How can lead be transformed into gold?
- On what basis are alternative forms of the periodic table constructed?
- Where would an element with an atomic number of 130 be placed in the modern periodic table, what properties would it have and how likely is it to be discovered?
- Why is the periodic table still a 'work in progress'?
- Why is it difficult to place hydrogen in the modern periodic table?
- Is it worthwhile finding any more new elements?
- Why aren't all the metals placed together in the periodic table?
- What makes some elements magnetic?
- Why do transition metals have multiple oxidation states?
- Is it an advantage or a disadvantage for elements to be unreactive?

Option 3: The lanthanoids and actinoids

Questions that may be explored in this investigation include:

- Why are some lanthanoids and actinoids so highly sought after?
- Can we live without lanthanoids and actinoids?
- Based on their usefulness for society, how would you compare the value of lanthanoids and actinoids with the value of other metal groups in the periodic table?
- Is it worth sending people to the Moon to mine for lanthanoids and actinoids?
- How are the lanthanoids and actinoids extracted from their ores?
- Do the lanthanoids and actinoids rust or corrode?
- Where are the lanthanoids and actinoids located in the periodic table and why have they been placed there?

Option 4: Using light to solve chemical puzzles

Questions that may be explored in this investigation include:

- What is a crystal, and why do crystals have regular faces?
- What makes synchrotron light useful?
- What can a synchrotron tell us about the differences between salt and sugar crystals?
- How does the composition of a crystal relate to the bonding within and the ratios of the elements present?
- Given that crystals are not alive or functioning, how is it that crystal structures are used to understand biological functions?
- What significant discoveries contributed to the development of X-ray crystallography as an analytical technique?
- How does cryoprotection preserve protein samples for analysis in a synchrotron?
- How has the use of a synchrotron enabled Nobel Prize winning research to occur?
- Why use synchrotron light to determine crystal structure when other sources of X-rays can be used?
- How does the IR beamline in the Australian Synchrotron enable the study of organic molecules and covalent bonding patterns?

Option 5: Glass

Questions that may be explored in this investigation include:

- What would life be like without glass? What might have to be used instead?
- How are the special properties achieved in particular forms of glass such as transition lenses, bullet-proof glass, safety glass, bendable glass, heat-proof glass, glass than can be switched on or off to become transparent or opaque, coloured glass and fibreglass?
- How are glass ornaments, glass jewellery and specialised laboratory glassware created?
- How does lightning make glass?
- Is glass safer to use than other forms of storage vessels?
- How useful is the type of glass that is produced as a by-product when iron is extracted from its ore in a blast furnace?

Option 6: Crude oil

Questions that may be explored in this investigation include:

- Do we need crude oil?
- Why does the composition of crude oil vary between different oil wells?
- How do different crude oil extraction methods compare with reference to ease of extraction and environmental impacts?
- What might we do if crude oil supplies run out?
- How does the time taken to produce crude oil compare with the time taken to use it?
- What fuels and other chemicals are derived from crude oil? How is crude oil processed to obtain them?
- Why are tar balls found on beaches after an oil spill?
- What are some of the issues surrounding society's demand for and use of crude oil? What strategies are being used to address some of these issues?

Option 7: Surfactants

Questions that may be explored in this investigation include:

- How do surfactants help clean up oil spills?
- Why is it so difficult to remove oil from bird feathers?
- Why can water birds drown if they still have detergent on their feathers?
- How are surfactants used in cooking, cosmetics and personal hygiene?
- How do hair shampoos differ from conditioners?
- How are soaps different from detergents?
- Why are different detergents made for cleaning different surfaces? How does their composition differ?
- How are surfactants designed for biodegradability?
- Does surfactant biodegradability affect performance?
- How can surfactants protect dams from drying out?
- What is the role of natural surfactants in the human body in breathing and digestion?

Option 8: Polymers and composite materials

Questions that may be explored in this investigation include:

- Are the biomaterials that replace body parts as effective as the original materials with reference to their properties and function?
- How can biomimicry help in developing new materials?
- How do new materials improve sporting performance?
- What are some of the new generation composite materials that have been designed to meet the emerging demands posed by space programs, and medical and technological developments?
- Should cars be made from shape memory metals?
- What makes some materials 'smart'?
- How do different types of radiation affect the structure and properties of polymer films?
- What applications for conductive polymers could be possible?
- What properties might a new 'super material' have, what might it be made from and what difference could it make to people's lives?
- Why would it be an advantage for polymers not to be biodegradable?
- How are new polymer and composite materials tested for safety?

Option 9: Nanomaterials

Questions that may be explored in this investigation include:

- What useful materials have resulted from nanomaterial research and what could be produced in future?
- How are nanochemists able to 'see' nanoparticles and manipulate atoms to build particular nanomaterials?
- Why do nanomaterials have different properties to their related macromaterials?
- Are nanomaterials safe?
- What difference can nanomaterials make to society and the environment?
- What can nanobots do?
- How are nanomaterials used in medicine and research?

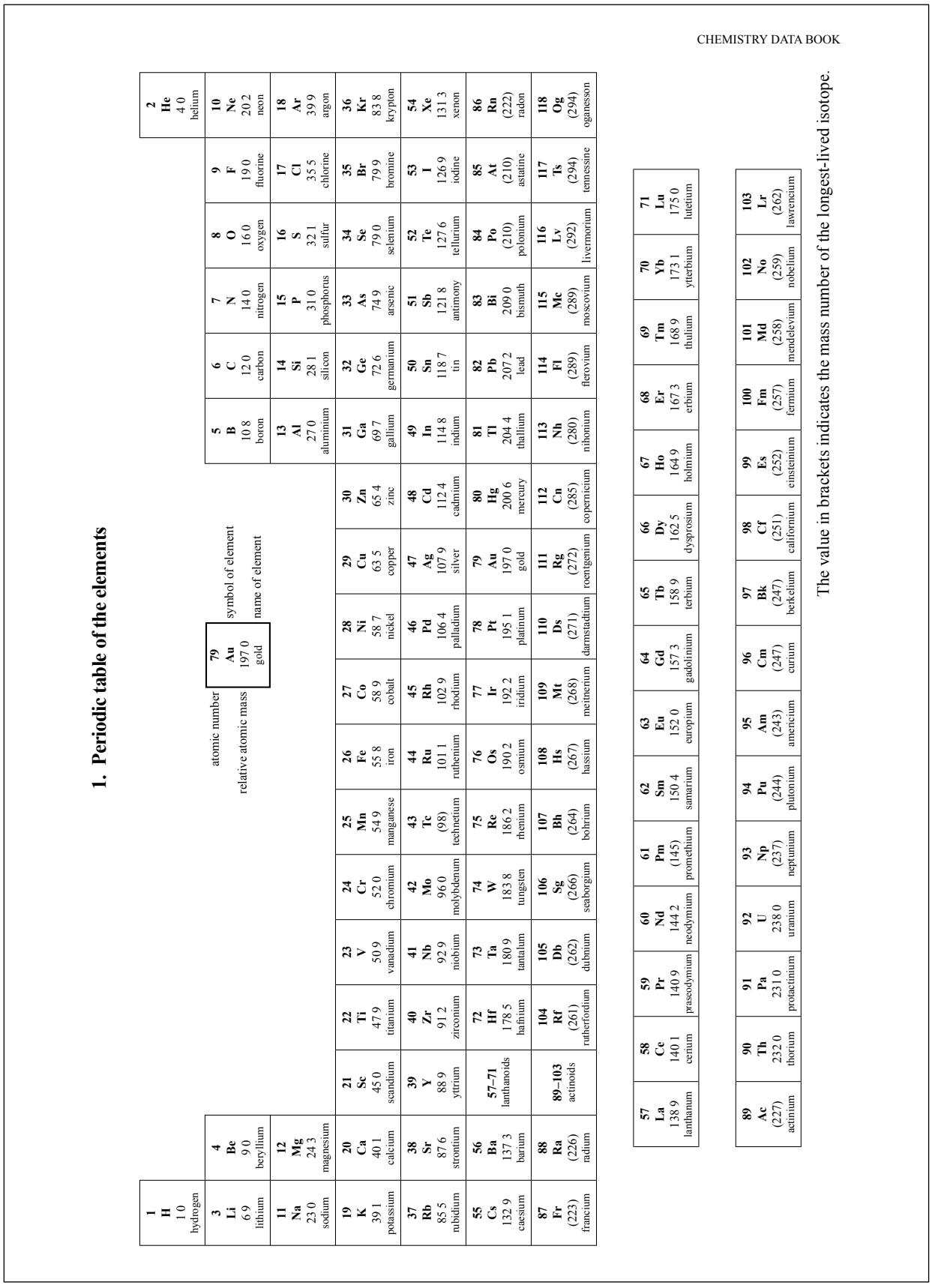
Option 10: The life cycle of a selected material or chemical

Questions that may be explored in this investigation include:

- What is the story behind the discovery of a selected material or chemical and why is it important?
- What is the relationship between the properties, structure and the nature and strength of the chemical bonding within the structure of a selected material or chemical?
- What are the main features of the life cycle of a selected material or chemical and how can these features be represented (they may include the extraction and processing of the raw materials; the manufacturing processes used, its packaging, distribution and use; any recycling and reuse; and its disposal at the end of its useful life)?
- What health and safety and environmental issues are involved in the manufacture, use and disposal of a selected material or chemical and how are these managed?
- What happens to a selected material or chemical at the end of its life?

VCAA DATA BOOK

For your Year 12 Chemistry exam, you will be given a data book that contains lots of useful information that you can use to help ace the exam. We thought it would be a good idea to get you familiar with the information that's in the data book, so we have put together a revised version that only includes things that are relevant to Year 11 Chemistry. Have fun!



CHEMISTRY DATA BOOK

1. Periodic table of the elements

																		2 **He** 4 0 helium
										5 **B** 10 8 boron	**6** **C** 12 0 carbon	**7** **N** 14 0 nitrogen	**8** **O** 16 0 oxygen	**9** **F** 19 0 fluorine	**10** **Ne** 20 2 neon
1 **H** 1 0 hydrogen										**13** **Al** 27 0 aluminium	**14** **Si** 28 1 silicon	**15** **P** 31 0 phosphorus	**16** **S** 32 1 sulfur	**17** **Cl** 35 5 chlorine	**18** **Ar** 39 9 argon
3 **Li** 6 9 lithium	**4** **Be** 9 0 beryllium									**31** **Ga** 69 7 gallium	**32** **Ge** 72 6 germanium	**33** **As** 74 9 arsenic	**34** **Se** 79 0 selenium	**35** **Br** 79 9 bromine	**36** **Kr** 83 8 krypton
11 **Na** 23 0 sodium	**12** **Mg** 24 3 magnesium									**49** **In** 114 8 indium	**50** **Sn** 118 7 tin	**51** **Sb** 121 8 antimony	**52** **Te** 127 6 tellurium	**53** **I** 126 9 iodine	**54** **Xe** 131 3 xenon
19 **K** 39 1 potassium	**20** **Ca** 40 1 calcium	**21** **Sc** 45 0 scandium	**22** **Ti** 47 9 titanium	**23** **V** 50 9 vanadium	**24** **Cr** 52 0 chromium	**25** **Mn** 54 9 manganese	**26** **Fe** 55 8 iron	**27** **Co** 58 9 cobalt	**28** **Ni** 58 7 nickel	**30** **Zn** 65 4 zinc	**31** **Ga** 69 7 gallium	**32** **Ge** 72 6 germanium	**33** **As** 74 9 arsenic	**34** **Se** 79 0 selenium	**35** **Br** 79 9 bromine	**36** **Kr** 83 8 krypton
37 **Rb** 85 5 rubidium	**38** **Sr** 87 6 strontium	**39** **Y** 88 9 yttrium	**40** **Zr** 91 2 zirconium	**41** **Nb** 92 9 niobium	**42** **Mo** 96 0 molybdenum	**43** **Tc** (98) technetium	**44** **Ru** 101 1 ruthenium	**45** **Rh** 102 9 rhodium	**46** **Pd** 106 4 palladium	**48** **Cd** 112 4 cadmium	**49** **In** 114 8 indium	**50** **Sn** 118 7 tin	**51** **Sb** 121 8 antimony	**52** **Te** 127 6 tellurium	**53** **I** 126 9 iodine	**54** **Xe** 131 3 xenon
55 **Cs** 132 9 caesium	**56** **Ba** 137 3 barium	**57–71** lanthanoids	**72** **Hf** 178 5 hafnium	**73** **Ta** 180 9 tantalum	**74** **W** 183 8 tungsten	**75** **Re** 186 2 rhenium	**76** **Os** 190 2 osmium	**77** **Ir** 192 2 iridium	**78** **Pt** 195 1 platinum	**79** **Au** 197 0 gold	**80** **Hg** 200 6 mercury	**81** **Tl** 204 4 thallium	**82** **Pb** 207 2 lead	**83** **Bi** 209 0 bismuth	**84** **Po** (210) polonium	**85** **At** (210) astatine	**86** **Rn** (222) radon
87 **Fr** (223) francium	**88** **Ra** (226) radium	**89–103** actinoids	**104** **Rf** (261) rutherfordium	**105** **Db** (262) dubnium	**106** **Sg** (266) seaborgium	**107** **Bh** (264) bohrium	**108** **Hs** (267) hassium	**109** **Mt** (268) meitnerium	**110** **Ds** (271) darmstadtium	**111** **Rg** (272) roentgenium	**112** **Cn** (285) copernicium	**113** **Nh** (280) nihonium	**114** **Fl** (289) flerovium	**115** **Mc** (289) moscovium	**116** **Lv** (292) livermorium	**117** **Ts** (294) tennessine	**118** **Og** (294) oganesson

57 **La** 138 9 lanthanum	**58** **Ce** 140 1 cerium	**59** **Pr** 140 9 praseodymium	**60** **Nd** 144 2 neodymium	**61** **Pm** (145) promethium	**62** **Sm** 150 4 samarium	**63** **Eu** 152 0 europium	**64** **Gd** 157 3 gadolinium	**65** **Tb** 158 9 terbium	**66** **Dy** 162 5 dysprosium	**67** **Ho** 164 9 holmium	**68** **Er** 167 3 erbium	**69** **Tm** 168 9 thulium	**70** **Yb** 173 1 ytterbium	**71** **Lu** 175 0 lutetium
89 **Ac** (227) actinium	**90** **Th** 232 0 thorium	**91** **Pa** 231 0 protactinium	**92** **U** 238 0 uranium	**93** **Np** (237) neptunium	**94** **Pu** (244) plutonium	**95** **Am** (243) americium	**96** **Cm** (247) curium	**97** **Bk** (247) berkelium	**98** **Cf** (251) californium	**99** **Es** (252) einsteinium	**100** **Fm** (257) fermium	**101** **Md** (258) mendelevium	**102** **No** (259) nobelium	**103** **Lr** (262) lawrencium

atomic number **79**
relative atomic mass **Au** 197 0
symbol of element
name of element
gold

The value in brackets indicates the mass number of the longest-lived isotope.

2. Electrochemical series

Reaction	Standard electrode potential (E^0) in volts at 25 °C
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77
$Au^+(aq) + e^- \rightleftharpoons Au(s)$	+1.68
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(1)$	+1.23
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g)$	+0.14
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	–0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	–0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	–0.25
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	–0.28
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	–0.40
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	–0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	–0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	–0.83
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	–1.18
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	–1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	–2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	–2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	–2.87
$K^+(aq) + e^- \rightleftharpoons K(s)$	–2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	–3.04

3. Chemical relationships

Name	Formula
number of moles of a substance	$n = \frac{m}{M}; \quad n = cV; \quad n = \frac{V}{V_m}$
heat energy released in the combustion of a fuel	$q = mc\Delta T$
electric charge	$Q = It$

4. Physical constants and standard values

specific Name	Symbol	Value
Avogadro constant	N_A or L	6.02×10^{23} mol^{-1}
heat capacity of water	c	4.18 kJ kg^{-1} K^{-1} or 4.18 J g^{-1} K^{-1}
density of water at 25 °C	d	997 kg m^{-3} or 0.997 g mL^{-1}

5. Unit conversions

Measured value	Conversion
1 litre (L)	1 dm^3 or 1×10^{-3} m^3 or 1×10^3 cm^3 or 1×10^3 mL

6. Metric (including SI) prefixes

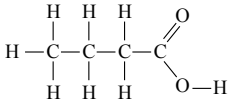
Metric (including SI) prefixes	Scientific notation	Multiplying factor
giga (G)	10^9	1 000 000 000
mega (M)	10^6	1 000 000
kilo (k)	10^3	1000
deci (d)	10^{-1}	0.1
centi (c)	10^{-2}	0.01
milli (m)	10^{-3}	0.001
micro (μ)	10^{-6}	0.000001
nano (n)	10^{-9}	0.000000001
pico (p)	10^{-12}	0.000000000001

7. Acid-base indicators

Name	pH range	Colour change from lower pH to higher pH in range
thymol blue (1st change)	1.2–2.8	red → yellow
methyl orange	3.1–4.4	red → yellow
bromophenol blue	3.0–4.6	yellow → blue
methyl red	4.4–6.2	red → yellow
bromothymol blue	6.0–7.6	yellow → blue
phenol red	6.8–8.4	yellow → red
thymol blue (2nd change)	8.0–9.6	yellow → blue
phenolphthalein	8.3–10.0	colourless → pink

8. Representations of organic molecules

The following table shows different representations of organic molecules, using butanoic acid as an example.

Formula	Representation
molecular formula	$C_4H_8O_2$
structural formula	
semi-structural (condensed) formula	$CH_3CH_2CH_2COOH$ or $CH_3(CH_2)_2COOH$
skeletal structure	