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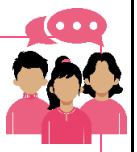
VCE Chemistry ¾
Introduction to Redox [1.6]
Workbook

Outline:

Learning Objectives:

- ☐ **CH34 [1.6.1]** - Apply oxidation numbers to find oxidant & reductant.
- ☐ **CH34 [1.6.2]** - Apply KOHES to write balanced half-equations and overall equations in acidic & basic conditions.

Section A: Introduction to Redox Reactions



Discussion: What are some examples of redox reactions?

What is a Redox Reaction?

Redox Reaction

- **Definition**: A redox reaction involves the _____ of _____.

REDOX	

Exploration: Redox Reaction

- Basic redox reaction (combustion of methane):

$$\mathbf{CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)}$$

- 'Half' reactions:

$$\mathbf{1.\ Oxidation:\ CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 8H^+(aq) + 8e^-}$$

$$\mathbf{2.\ Reduction:\ 2O_2(g) + 8H^+(aq) + 8e^- \rightarrow 4H_2O(l)}$$

- Electrons in the first oxidation reaction: [lost] / [gained]
- Electrons in the second reduction reaction: [lost] / [gained]
- What happens to the electrons from equation 1 to equation 2? _____

Discussion: Can the oxygen (O_2) reaction occurs by itself without the methane (CH_4) reaction occurring?

[Yes] / [No]

Discussion: Can the methane (CH_4) reaction occurs by itself without the oxygen (O_2) reaction occurring?

[Yes] / [No]

Key Takeaways

- ☑ Redox reactions must occur in pairs.
- ☑ The two half-equations cannot occur by themselves and must occur together.
- ☑ Redox is the exchange of electrons – one substance gives away electrons, and the other substance takes in electrons.

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Section B: Redox Reactions & Oxidation Numbers

Sub-Section: Oxidation and Reduction Reactions

What do oxidation and reduction mean?

Oxidation and Reduction Reactions

- Oxidation reactions always result in the _____ of electrons.
- Reduction reactions always result in the _____ of electrons.

How do we remember this?

TIP:

OIL RIG

Why is this the case and what does it mean if something is oxidised?

Oxidation Number

- **Definition**: States the hypothetical _____ of an atom if the bonding is purely ionic.
- According to the **Octet Rule**, atoms generally want _____ electrons in its outer shell.

Exploration: Sodium (Na)

- Consider sodium (Na):

Na

- **Key Information:**

Number of Protons	Number of Electrons	Overall Charge	Oxidation Number

- **Electron Configuration:** _____
- What will the sodium (Na) do to obtain a full outer shell? _____

Exploration: Sodium Ion (Na^+)

- Sodium usually exists in the following form:

Na^+

- **Key Information:**

Number of Protons	Number of Electrons	Overall Charge	Oxidation Number

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What happened? How does this link to oxidation numbers?

Exploration: Overall Reaction

- Complete the equation including electrons. *(Label Below)*

$$Na \rightarrow Na^{+}$$

- Electrons: [gained] / [lost]
- Reaction: [reduction] / [oxidation]
- Oxidation Number:

Before: Na	After: Na^{+}

- Oxidation Number: [increases] / [decreases]
- Substance becomes: [more] / [less] oxidised
- Electrons: [gained] / [lost]
- Substance's Overall Charge: [increases] / [decreases]

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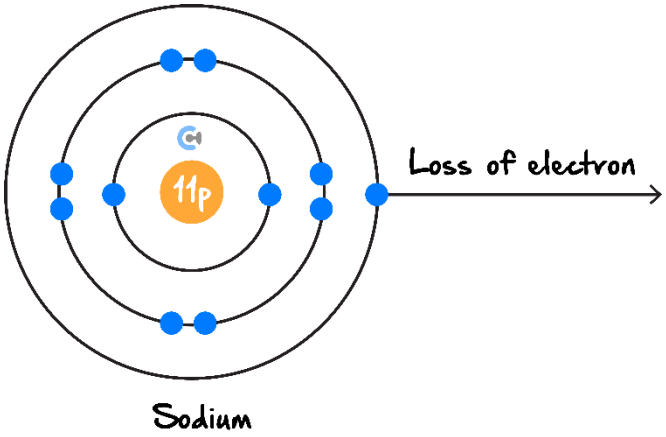
Misconception

"In the following reaction, electrons are produced, and thus, it is a reduction reaction as we gain electrons!"

$$Na \rightarrow Na^+ + e^-$$

TRUTH: Electrons are lost in this reaction!

- Think about it from the point of view of the sodium itself:



What if we want to go the other way around?

Exploration: Opposite Reaction

- What if we want to go from a Na^+ ion to Na? *(Label Below)*

$$Na^+ + e^- \rightarrow Na$$

- Oxidation number: [increases] / [decreases] / [no change]
- What happens to the sodium compared to before? [More] / [Less] oxidised
- Type of Reaction: [reduction] / [oxidation]
- Electrons: [gained] / [lost]

TIP: Think of oxidation numbers and tell us how oxidised the substance is. The higher the oxidation number, the more oxidised the substance is!

Reduction & Oxidation Reactions

Oxidation Reaction	Reduction Reaction
Electrons are **[Gained] / [Lost].**	**Electrons** are **[Gained] / [Lost].**
Oxidation Number **[Increases] / [Decreases].**	Oxidation Number **[Increases] / [Decreases].**

Extension: **Which one is more likely to occur in nature? Na^+ reducing into Na or Na oxidising into Na^+ ? (Hint: Think back to the Octet Rule)**

$$Na \rightarrow Na^+ + e^-$$

$$Na^+ + e^- \rightarrow Na$$

NOTE: We'll cover the 'likelihood' for certain reactions in the next booklet.

Question 1

Which of the following statements is true for an oxidation reaction?

A. Electrons are gained by the species.

B. Electrons are lost by the species.

C. The oxidation state decreases.

D. Electrons are written on the left side of the equation.

Question 2

Which of the following represents a reduction in half-reaction?

A. $Fe \rightarrow Fe^{3+} + 3e^-$

B. $Cl_2 + 2e^- \rightarrow 2Cl^-$

C. $Cu \rightarrow Cu^{2+} + 2e^-$

D. $Na \rightarrow Na^+ + e^-$

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Sub-Section: Calculating Oxidation Numbers

Context

- When calculating oxidation numbers, there are a couple of **rules**.

Exploration: Oxidation Number of Elements

Rule #1: Free Elements

- **Oxidation Number:**

Na or F

- **Oxidation Number:**

$$Cl_2 \text{ or } O_3$$

- **Rule #1:** The oxidation number of any free element (exists by itself) is always ______________.

Rule #2: Ions

- **Oxidation Number:**

$$Na^+ \text{ or } F^-$$

- **Rule #2:** The oxidation number of any ion is just the __________________ of the ion.

Rule #3: Ionic Compound

- **Oxidation Number:**

NaF

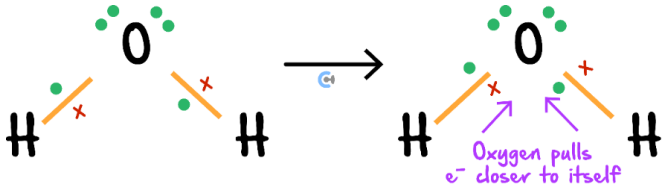
- **Rule #3:** The oxidation number in ______________ compounds is the same as the charge carried by the cation and anion respectively.

NOTE: This rule only works for ionic bonding it does not work for molecules which are covalently bonded to each other. Instead for covalent bonding, a different method must be used.

What about covalent molecules?

Exploration: Oxidation Number of Molecules

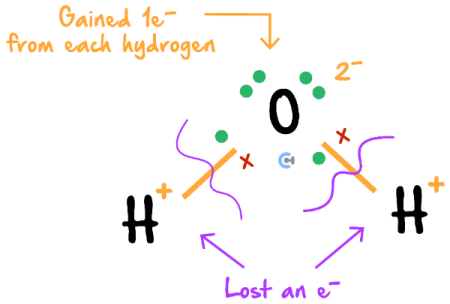
- Consider water (H_2O):



- Which is more **electronegative**? [Hydrogen] / [Oxygen] atoms

- **Rule**: When calculating oxidation numbers, visualise all bonds such as these covalent bonds as being _____ bonds!

- Find the **hypothetical charge** on each atom if it were to be an ionic bond.



Effective Charge on Hydrogen Atom (Oxidation Number)	Effective Charge on Oxygen Atom (Oxidation Number)

- **Each Hydrogen** is typically going to have an oxidation number of $+1$.

- **Oxygen** is typically going to have an oxidation number of -2.

Exploration: Hydrogen Oxidation Number

With Non-Metals (Majority)

- Oxidation number of hydrogen (H) when combined with a non-metal: _____
- Electronegativity of non-metals compared to hydrogen: [stronger] / [weaker]
- Electrons pull relative to hydrogen: Pulls [towards] / [away from] hydrogen
- **Oxidation Number:**

HCl

CH_4

- **Rule**: The _____ of all the oxidation numbers in a neutral compound is _____.

With Metals (Minority)

- Oxidation number of hydrogen (H) when it is combined with a metal is _____.
- Electronegativity of non-metals compared to hydrogen: [stronger] / [weaker]
- Electrons pull relative to hydrogen: Pulls [towards] / [away from] hydrogen
- **Oxidation Number:**

NaH or CaH_2

Oxidation Number Rules

- Isolated Elements (e.g., H_2):
- Ions (e.g., Na^+):
- Ionic Compounds (e.g., $NaCl$):
- Oxygen (O):
- Hydrogen (H):
- Sum of oxidation numbers in the compound is equal to (e.g., H_2SO_4 or MnO_4^-):

Let's look at a question together!

Question 3 Walkthrough.

Find the oxidation number of sulphur in sulphuric acid.

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Your Turn!

Question 4

Find the oxidation number for all elements in each of the following molecules.

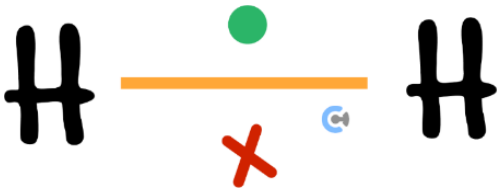
a. H_2O	**c.** NaH
b. NH_3	**d.** H_2

Misconception

The oxidation number of $\mathbf{H_2}$ *is* $+1$ *or* -1.

TRUTH:

- Consider hydrogen gas (H_2):



- Which hydrogen is more electronegative? _____
- What is the 'effective charge' of each of them? _____

NOTE: The one exception is when hydrogen exists as H_2 - oxidation number of hydrogen is just 0.

ALSO NOTE: The oxidation number of hydrogen is usually just $+1$.

Active Recall: What is the usual oxidation number of oxygen?

Exploration: Oxygen Oxidation Number

- Oxidation Number:

$$NO_2 \text{ or } SO_3$$

- The **sum** of oxidation numbers in a **polyatomic** ion is equal to its **charge**.

$$CO_3^{2-} \text{ or } Cr_2O_7^{2-}$$

NOTE: The oxidation number of oxygen is not -2 when the compound contains $O-O$ bonds such as in: $O_2, O_3, H_2O_2, O_2^{2-}$ and so on, but this is also rarely tested in VCAA.

Let's look at a question together!

Question 5 Walkthrough.

Find the oxidation number of carbon in citrate.

Your Turn!

Question 6

State the oxidation number for the element specified in the molecule/ion provided.

a. Oxidation number of sulphur in sulphate.

b. Oxidation number of phosphorous in phosphate.

c. Oxidation number of nitrogen in nitrate.

d. Oxidation number of chromium in dichromate.

Question 7

Determine the oxidation number of chlorine in each of the following:

a. HClO

_____

b. $NaClO_4$

_____

c. ClO_3^-

_____

Question 8 Additional Question.

Vanadium is an example of a transition metal that has more than one oxidation state.

The following ions or compounds containing vanadium are known to exist:

V_2O_5 VO_2^+ VO^{2+} VCl_3 VCl_2

The oxidation number of vanadium in the above compounds is:

A. -1 and -2.

B. $+2$ and $+3$ only.

C. $+3$ and $+5$ only.

D. $+2, +3, +4$ and $+5$.

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Sub-Section: Oxidants and Reductants

REMINDER: Don't forget that oxidation is the increase in oxidation number (loss of electrons), and reduction is the decrease in oxidation number (gain of electrons).

Exploration: Oxidising Agents and Reducing Agents

1. Oxidation: $CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 8H^+(aq) + 8e^-$

2. Reduction: $2O_2(g) + 8H^+(aq) + 8e^- \rightarrow 4H_2O(l)$

- Oxidising agent (oxidant):
  - **Purpose:** _____.
  - Type of reaction oxidant undergoes: [reduction] / [oxidation]
- Reducing agent (reductant):
  - Reducing agent, abbreviated to 'reductant', causes [reduction] / [oxidation] to the other species.
  - Type of reaction reductant undergoes: [reduction] / [oxidation]

TIP: Just think of the oxidising agent/oxidant causing oxidation to the other species.

Oxidising Agents and Reducing Agents

- **Oxidant**: Causes oxidation to other species, itself undergoes reduction.
- **Reductant**: Causes reduction to other species, itself undergoes oxidation.

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Try a Question!

Question 9 Walkthrough.

In the following reaction:

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$

a. Which species has been oxidised and which species has been reduced?

_____

_____

b. Identify the:

i. Oxidant.

_____

ii. Reductant.

_____

Discussion: The oxidation number of oxygen (O) does not change. Has the oxygen atom been oxidised or reduced?

Your Turn!

Question 10 (1 mark)

In the following reaction:

$$\mathbf{Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)}$$

a. Which species has been oxidised and which species has been reduced? (0.5 marks)

_____

_____

b. Identify the: (0.5 marks)

Oxidant: _____ Reductant: _____

Question 11 (1 mark)

In the reaction:

$$CuO(s) + Zn(s) \rightarrow ZnO(s) + Cu(s)$$

a. State which species has been oxidised. (0.5 marks)

_____

b. State the oxidising agent. (0.5 marks)

_____

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Question 12 (1 mark)

In the reaction:

$$HOCl(aq) + H^+(aq) + 2I^-(aq) \rightarrow I_2(aq) + Cl^-(aq) + H_2O(l)$$

A. Hydrogen is reduced and oxygen oxidised.

B. Chlorine is reduced and iodine oxidised.

C. Iodine is reduced and hydrogen oxidised.

D. Chlorine is oxidised and oxygen reduced.

Question 13 (1 mark)

In which of the following reactions does the metal atom show the greatest change in oxidation state?

A. MnO_4^- to Mn^{2+}

B. MnO_2 to $Mn(OH)_3$

C. PbO_2 to $PbSO_4$

D. VO_2^+ to VO^{2+}

Question 14 (1 mark)

Which one of the following is least likely to be a product of a redox reaction between sulphuric acid (H_2SO_4) and zinc metal (Zn)?

A. H_2

B. H_2S

C. SO_2

D. SO_3

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Sub-Section: Conjugate Redox Pairs

Exploration: Conjugate Redox Pairs

- Oxidant and Reductant in Equation: *(Label Below)*

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$

- Oxidant and Reductant in Flipped Equation: *(Label Below)*

$$Cu(s) + H_2O(l) \longrightarrow CuO(s) + H_2(g)$$

- Products:

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$$

- **Representation**: Always shows the oxidising agent/oxidant first.
- Conjugate Redox Pairs for the above reaction:

Let's look at a question together!

Question 15 Walkthrough.

Sulphur can exist in the form of thiosulphate, and can react and form $S_2O_6^{2-}$. Write the conjugate redox pair for the reaction.

_____

_____

Your Turn!

Question 16

The following reaction turns bromine into bromate ions.

$$Br_2(l) + 6H_2O(l) \rightarrow 2BrO_3^-(aq) + 12H^+(aq) + 10e^-$$

Write the conjugate redox pair.

_____

_____

Question 17

Write the conjugate redox pair between permanganate ions and manganese dioxide.

_____

_____

Key Takeaways

Oxidation Reaction	Reduction Reaction
Electrons are Lost.	Electrons are Gained.
Oxidation Number Increases.	Oxidation Number Decreases.

- **Oxidation Number Rules**
  - Isolated Elements (e.g., H_2): 0
  - Ions (e.g., Na^+): Charge
  - Ionic Compounds (e.g., $NaCl$): Charge
  - Oxygen (O): -2
  - Hydrogen (H): $+1$
  - The sum of oxidation numbers in compound is equal to (e.g., H_2SO_4 or MnO_4^-): Charge
- **Oxidant**: Causes oxidation to other species, itself undergoes reduction.
- **Reductant**: Causes reduction to other species, itself undergoes oxidation.
- In conjugate redox pairs, the oxidant is always written first.

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Section C: Balancing Redox Half Equations

Sub-Section: Balancing Equations in Acidic (H^+) Environments

Context

- How do we actually balance these half-equations?
- How can the following reaction be balanced?

$$K \rightarrow K^+$$

  - Balance the _____.

- How can the following equation be balanced?

$$Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$$

- The equation is balanced by **introducing other reactants/products** which are available.

Exploration: Balancing Equation

- Where does this reaction take place? (Hint: Look at the states.) _____
- Reagents Available: _____
- Other Substances present in Water: _____
- Equation balanced by using _____and _____ as the reagents.

Exploration: Acidic Environments

- Reagents available in acidic environments:

$$H_2O \ \& \ H^+$$

- **Balancing Oxygen:**
  - Substance which contains oxygen: _____
  - Oxygen is balanced using: _____

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$$

  - What is unbalanced now? _____
- **Balancing Hydrogen:**
  - Hydrogen is balanced using: _____

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

  - What is unbalanced now? _____
- **Balancing Charge:**
  - Charge is balanced using: _____

$$14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

Definition

Balancing Equations in Acidic Conditions

- Balancing Equation Steps:
  - Balanced the _____
  - Balanced the _____ by adding _____
  - Balanced the _____ by adding _____
  - Balanced the _____ by adding _____
  - Included the _____
- Acronym: _____

Let's try a question together!

Question 18 Walkthrough.

In acidic conditions, nitrate (NO_3^-) can react and turn into nitrogen dioxide gas (NO_2).

a. Write a balanced half-equation depicting this process.

_____

_____

_____

b. Is this an oxidation or reduction reaction?

_____

c. Is nitrate (NO_3^-) an oxidant or reductant?

_____

Recall!

Active Recall: What does KOHES stand for?

- K: _____
- O: _____
- H: _____
- E: _____
- S: _____

Your Turn!

Question 19

Complete the balanced half-equations, and state whether it is a reduction or oxidation reaction.

a. Copper (II) ions (Cu^{2+}) turning into copper solid (Cu).

_____

_____

Type of Reaction: [Reduction] / [Oxidation]

b. Manganese metal (Mn) turning into permanganate (MnO_4^-).

_____

_____

Type of Reaction: [Reduction] / [Oxidation]

c. Dihydrogen phosphate ($H_2PO_4^-$) turning into a phosphorous solid (P).

_____

_____

Type of Reaction: [Reduction] / [Oxidation]

d. Chlorine gas (Cl_2) turning into chlorate ions (ClO_3^-).

_____

_____

Type of Reaction: [Reduction] / [Oxidation]

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Question 20

In the following redox reactions, extract the oxidation and reduction reactions, and identify the oxidant and reductant.

a. $Cr_2O_7^{2-}(aq) + 8H^+(aq) + 3SO_3^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq) + 4H_2O(aq)$

Reduction Equation: _____

Oxidation Equation: _____

Oxidant: _____ Reductant: _____

b. $6MnO_4^-(aq) + 18H^+(aq) + 5S_2O_3^{2-}(aq) \rightarrow 6Mn^{2+}(aq) + 5S_2O_6^{2-}(aq) + 9H_2O(l)$

Reduction Equation: _____

Oxidation Equation: _____

Oxidant: _____ Reductant: _____

Question 21 Additional Question.

Write the oxidation and reduction reaction from the following reaction, and identify the oxidants and reductants.

$$2MnO_4^-(aq) + 3C_2O_4^{2-}(aq) + 8H^+(aq) \rightarrow 2MnO_2(aq) + 4H_2O(l) + 6CO_2(g)$$

Reduction Equation: _____

Oxidation Equation: _____

Oxidant Equation: _____ Reductant: _____

Question 22 Additional Question.

Redox reactions occur in the human body as well as in electrochemical cells.

Nicotinamide adenine dinucleotide (NAD) is a vital coenzyme for energy production in the human body. It exists in two forms: an oxidised form, NAD^+, and a reduced form, NADH.
NAD is involved in the conversion of ethanol, CH_3CH_2OH, to ethanal, CH_3CHO, in the human body. The overall equation for this redox reaction is:

$$CH_3CH_2OH + NAD^+ \rightarrow CH_3CHO + NADH + H^+$$

a. Write the two half-equations for this redox reaction. States are not required.

Oxidation half-equation _____

Reduction half-equation _____

b. Identify the reducing agent in this redox reaction.

_____

Exploration: Trick to double-check if the equation is balanced

- Consider the following equation:

$$\mathbf{MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O}$$

- **Oxidation Numbers of Manganese:**

$\mathbf{MnO_4^-(aq)}$	$\rightarrow$	$\mathbf{Mn^{2+}(aq)}$

- Change in Oxidation Number: _____
- Number of electrons gained/lost: _____ electrons [gained] / [lost]
- **Conclusion**: In redox reactions, the change in oxidation number is caused by the substance gaining/losing electrons.

Exploration: Checking this Trick I

- Consider the following equation:

$$SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$$

- Oxidation Numbers of Sulphur:

$SO_3^{2-}(aq)$	$\rightarrow$	$SO_4^{2-}(aq)$

- Change in Oxidation Number: _____
- Number of electrons gained/lost: _____ electrons [gained] / [lost]

Exploration: Checking this Trick II

- Consider the following equation:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

- Oxidation Numbers of Chromium:

$Cr_2O_7^{2-}(aq)$	$\rightarrow$	$2Cr^{3+}(aq)$

- Change in Oxidation Number: _____
- Number of electrons gained/lost: _____ electrons [gained] / [lost]
- Does this match up? Why? How? [Yes] / [No]

NOTE: As there are **two Cr** atoms, it needs **double** the amount of electrons!

Double Checking Electrons align with the Equation

- The number of electrons lost/gained should align with the change in oxidation number.

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Sub-Section: Forming the Overall Equation

We can now get the two half-equations, but how do we form an overall redox reaction?

Exploration: Forming the Overall Equation

- What happens to electrons in the overall reaction? *(Label Below)*

$$\textbf{Oxidation: } K \rightarrow K^+ + e^-$$

$$\textbf{Reduction: } 2H^+ + 2e^- \rightarrow H_2$$

- Are electrons present in the overall reaction? [Yes] / [No]
- Combine the two half-reactions by _____ the _____.

Forming the Overall Equation

- Cancel out electrons by finding the lowest common multiple.

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Let's try a question together!

Question 23 Walkthrough.

Write the overall reaction given the half-equations below:

Oxidation:

$$K \rightarrow K^+ + e^-$$

Reduction:

$$2H^+ + 2e^- \rightarrow H_2$$

Overall:

_____

_____

Your Turn!

Question 24

Iron (III) Fe^{3+} can be formed from iron (II) Fe^{2+} when reacted with dichromate ions ($Cr_2O_7^{2-}$). Cr^{3+} ions are also formed in the process. Here are the two equations:

1. Oxidation: $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

2. Reduction: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$

Write the balanced overall reaction.

_____

_____

_____

Question 25 (5 marks)

Sulphur dioxide (SO_2) is a chemical of major industrial significance.

a. SO_2 gas can be produced in a reaction between concentrated sulphuric acid (H_2SO_4) and sodium metal. A solution containing Na^+ ions are also formed.

Write balanced equations for the:

i. Oxidation reaction. (1 mark)

_____

ii. Reduction reaction. (1 mark)

_____

_____

iii. Overall reaction, showing the states of all reactants and products. (2 marks)

_____

_____

_____

SO_2 can also be produced in a chemical reaction between zinc sulphide ($ZnSO_3$) and hydrochloric acid according to the equation:

$$ZnSO_3(s) + 2\,H^+(aq) \rightarrow Zn^{2+}(aq) + SO_2(g) + H_2O(l)$$

b. Is this reaction also a redox reaction? Explain your answer. (1 mark)

_____

_____

NOTE: When writing an overall redox reaction, you need to first balance the two half-equations first!

Question 26 Additional Question.

Balance the following equation:

$$Zn(s) + MnO_4^-(aq) \rightarrow Zn^{2+}(aq) + Mn^{2+}$$

Question 27 Additional Question.

Balance the following equation:

$$Sn^{2+}(aq) + IO_3^-(aq) \rightarrow Sn^{4+}(aq) + I^-(s)$$

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Sub-Section: Balancing Equations in Basic/Alkaline (OH^-) Environments

What happens when balancing equations in basic conditions?

Exploration: Basic Conditions

- Consider the following equation in **basic conditions**:

$$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \longrightarrow SO_2(g) + 2H_2O(l)$$

- Can H^+ be used to balance the equation? [Yes] / [No]
- What should be used instead? _____
- **Reactants used in Basic Conditions:**

$$H_2O \ \& \ OH^-$$

Method to Balance in Basic Conditions

1. Balance in _____ conditions first using KOHES.
2. _____ hydrogen ions (H^+) by adding _____

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Let's look at a question together!

Question 28 Walkthrough.

Rewrite the following equation in basic conditions:

$$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \longrightarrow SO_2(g) + 2H_2O(l)$$

Question 29 Walkthrough.

Write a balanced equation in basic conditions for the reaction where manganese ions (Mn^{2+}) turns into permanganate ions (MnO_4^-).

TIP: You can also **double-check** whether the half equation has been balanced by **counting the charge** on both sides of the equation.

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Your Turn!

Question 30

Write the reaction of $Cr_2O_7^{2-}$ ions to Cr^{3+} ions under alkaline conditions.

Question 31

Write the reaction of IO_3^- into I_2 under alkaline conditions.

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Question 32

Write the reaction of O_2 into H_2O under alkaline conditions.

Question 33 Additional Question.

Balance the following unbalanced overall equations in basic conditions:

$$S^{2-} + NO_3^- \rightarrow S + NO$$

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Contour Check

Learning Objective: [1.6.1] - Apply oxidation numbers to find oxidant & reductant.

Study Design

Redox reactions as simultaneous oxidation and reduction processes, and the use of oxidation numbers to identify the reducing agent, oxidising agent and conjugate redox pairs.

Key Takeaways

- ☐ Redox reactions must occur _____.
- ☐ Redox is the _____ of electrons - one substance gives away electrons, the other substance takes in electrons.

Oxidation Reaction	Reduction Reaction
Electrons are **[Gained] / [Lost].**	**Electrons** are **[Gained] / [Lost].**
Oxidation Number **[Increases] / [Decreases].**	Oxidation Number **[Increases] / [Decreases].**

- ☐ **Oxidation Number Rules**
  - ○ Isolated Elements (e.g., H_2):
  - ○ Ions (e.g., Na^+):
  - ○ Ionic Compounds (e.g., $NaCl$):
  - ○ Oxygen (O):
  - ○ Hydrogen (H):
  - ○ Sum of oxidation numbers in compound is equal to (e.g., H_2SO_4 or MnO_4^-):
- ☐ **Oxidant**: Causes [reduction] / [oxidation] to other species, itself undergoes [reduction] / [oxidation].

- **Reductant**: Causes [reduction] / [oxidation] to other species, itself undergoes [reduction] / [oxidation].
- In conjugate redox pairs, the [oxidant] / [reductant] is always written first.

Learning Objective: [1.6.2] - Apply KOHES to write balanced half-equations and overall equations in acidic & basic conditions.

Study Design

The writing of balanced half-equations (including states) for oxidation and reduction reactions, and the overall redox cell reaction in both acidic and basic conditions.

Key Takeaways

- Balancing Equation Steps:
  - Balanced the _____.
  - Balanced the _____ by adding _____.
  - Balanced the _____ by adding _____.
  - Balanced the _____ by adding _____.
  - Included the _____.
- Acronym: _____
- Balancing in Basic Conditions:
  1. Balance in _____ conditions first using KOHES.
  2. _____ hydrogen ions (H^+) by adding _____.
- Number of electrons lost/gained should align with change in _____.
- Forming Overall Equation: Cancel out _____ by finding _____.

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