



VCE CHEMISTRY 2009

YEAR 12 TRIAL EXAM UNIT 4

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Time allowed: 90 minutes

Total marks: 82

20 Multiple Choice Questions

6 Short Answer Questions

An Answer Sheet is provided for Section A.

Answer all questions in Section B in the space provided.

To download the Chemistry Data Book please visit the VCAA website:

http://www.vcaa.vic.edu.au/vce/studies/chemistry/chem1_sample_2008.pdf Page 20

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Student Name.....

VCE Chemistry 2009 Year 12 Trial Exam Unit 4

Student Answer Sheet

Instructions for completing test. Use only a 2B pencil. If you make a mistake erase and enter the correct answer. Marks will not be deducted for incorrect answers.

Write your answers to the Short Answer Section in the space provided directly below the question. There are 20 Multiple Choice questions to be answered by circling the correct letter in the table below.

<i>Question 1</i>	A	B	C	D	<i>Question 2</i>	A	B	C	D
<i>Question 3</i>	A	B	C	D	<i>Question 4</i>	A	B	C	D
<i>Question 5</i>	A	B	C	D	<i>Question 6</i>	A	B	C	D
<i>Question 7</i>	A	B	C	D	<i>Question 8</i>	A	B	C	D
<i>Question 9</i>	A	B	C	D	<i>Question 10</i>	A	B	C	D
<i>Question 11</i>	A	B	C	D	<i>Question 12</i>	A	B	C	D
<i>Question 13</i>	A	B	C	D	<i>Question 14</i>	A	B	C	D
<i>Question 15</i>	A	B	C	D	<i>Question 16</i>	A	B	C	D
<i>Question 17</i>	A	B	C	D	<i>Question 18</i>	A	B	C	D
<i>Question 19</i>	A	B	C	D	<i>Question 20</i>	A	B	C	D

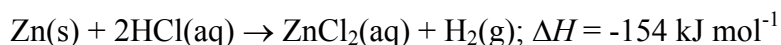
VCE Chemistry 2009 Year 12 Trial Exam Unit 4

Multiple Choice Questions – Section A

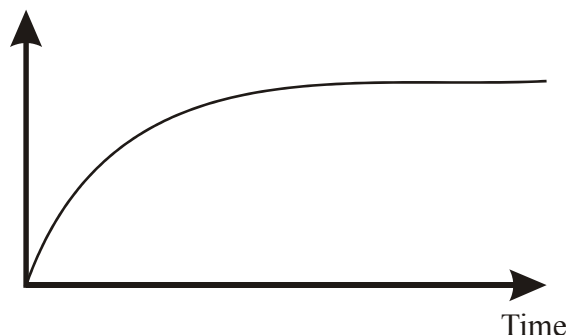
*Section A consists of 20 multiple-choice questions.
Section A is worth approximately 25 per cent of the marks available.
Choose the response that is **correct** or **best answers** the question.
Indicate your choice on the answer sheet provided.*

Question 1

The reaction between zinc and hydrochloric acid may be represented by the equation



In an investigation of this reaction using excess Zn and 2 M HCl(aq) in an open flask, the following graph was plotted from data collected.



Which of the following would be a suitable quantity for the vertical axis of the graph?

- A. Number of collisions per second.
- B. Rate of reaction.
- C. Temperature.
- D. Number of ions.

Question 2

Zinc-air batteries used in hearing aids and experimental electrical vehicles have high energy density and are relatively inexpensive to produce.

The overall redox reaction for a zinc-air battery delivering electrical energy may be described by the equation



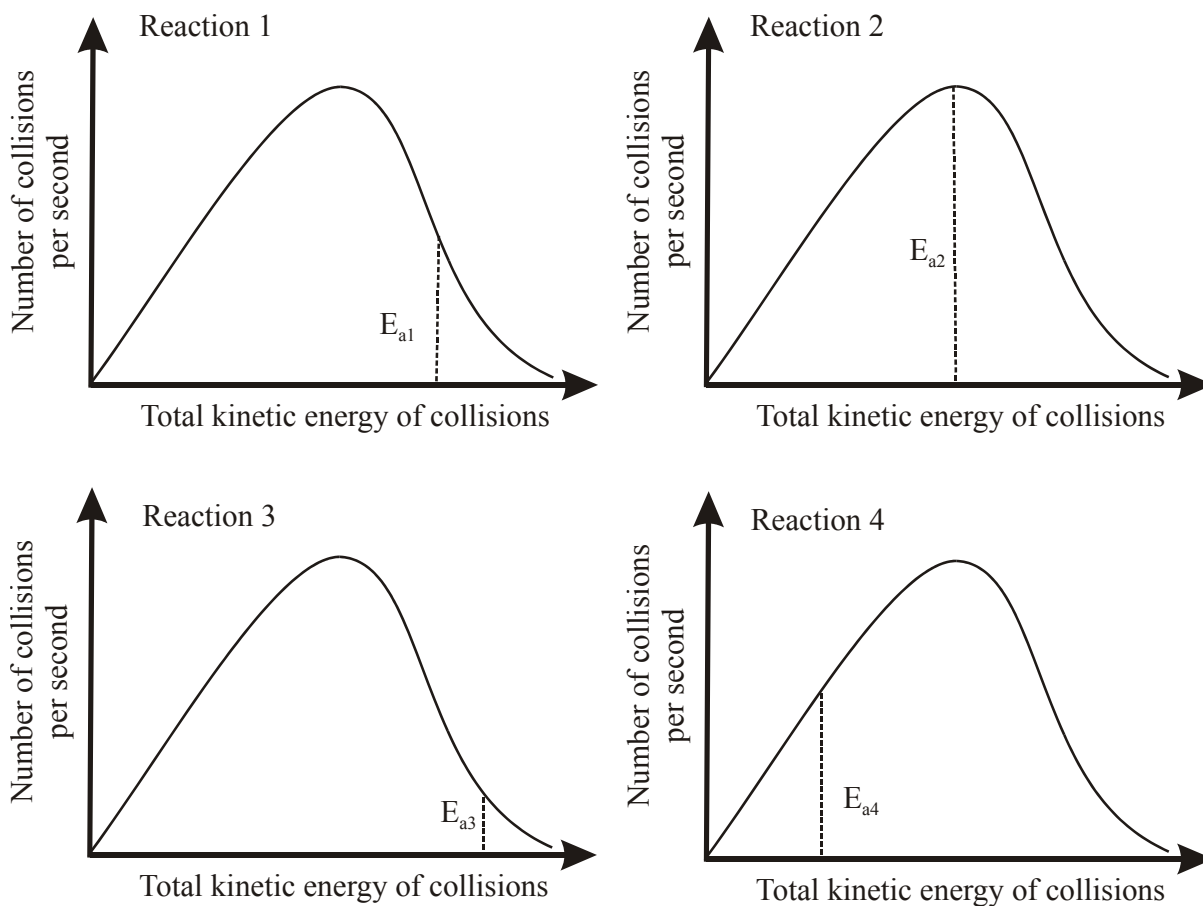
In this battery, zinc, Zn

- A. forms the positive electrode, and is reduced.
- B. forms the positive electrode, and is oxidised.
- C. forms the negative electrode, and is reduced.
- D. forms the negative electrode, and is oxidised.

Question 3

Shown below are kinetic energy distribution curves for four different reactions, all occurring at the same temperature.

The activation energies of the four reactions are represented by the symbols E_{a1} , E_{a2} , E_{a3} and E_{a4} respectively.

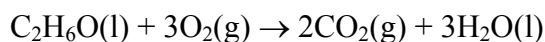
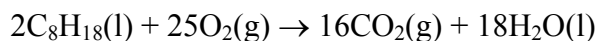


Which reaction is most likely to have the fastest reaction rate?

- A. Reaction 1.
- B. Reaction 2.
- C. Reaction 3.
- D. Reaction 4.

Question 4

The use of E10 fuel in car engines produces carbon dioxide due to the combustion of octane and ethanol according to the equations



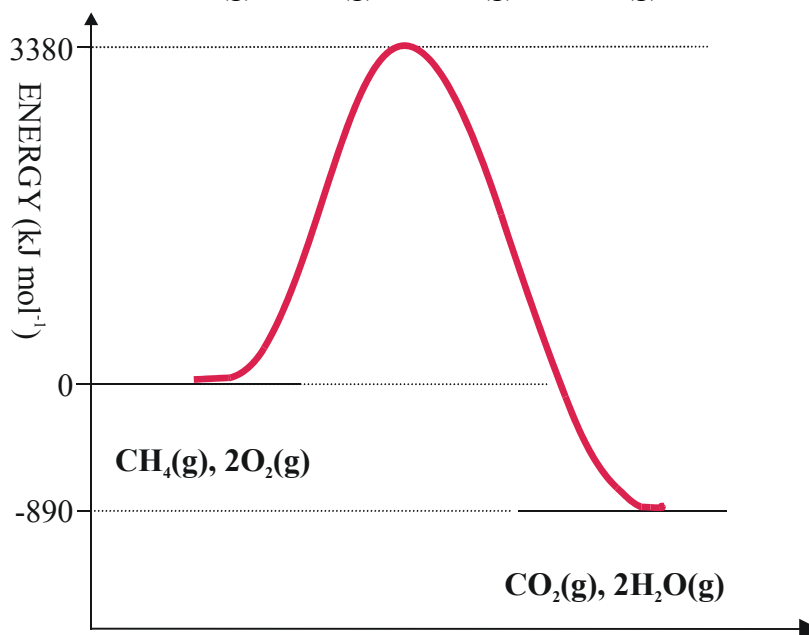
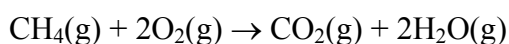
In a comparison of the complete combustion of samples of pure octane and pure ethanol, both samples release 2000 kJ of energy.

The ratio $n(\text{CO}_2)$ produced by octane sample : $n(\text{CO}_2)$ produced by ethanol sample is closest to

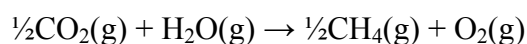
- A. 1 : 1
- B. 1 : 4
- C. 4 : 1
- D. 8 : 1

Question 5

Shown below is the energy profile for combustion of methane according to:



According to the information provided on this profile, the activation energy for the reaction

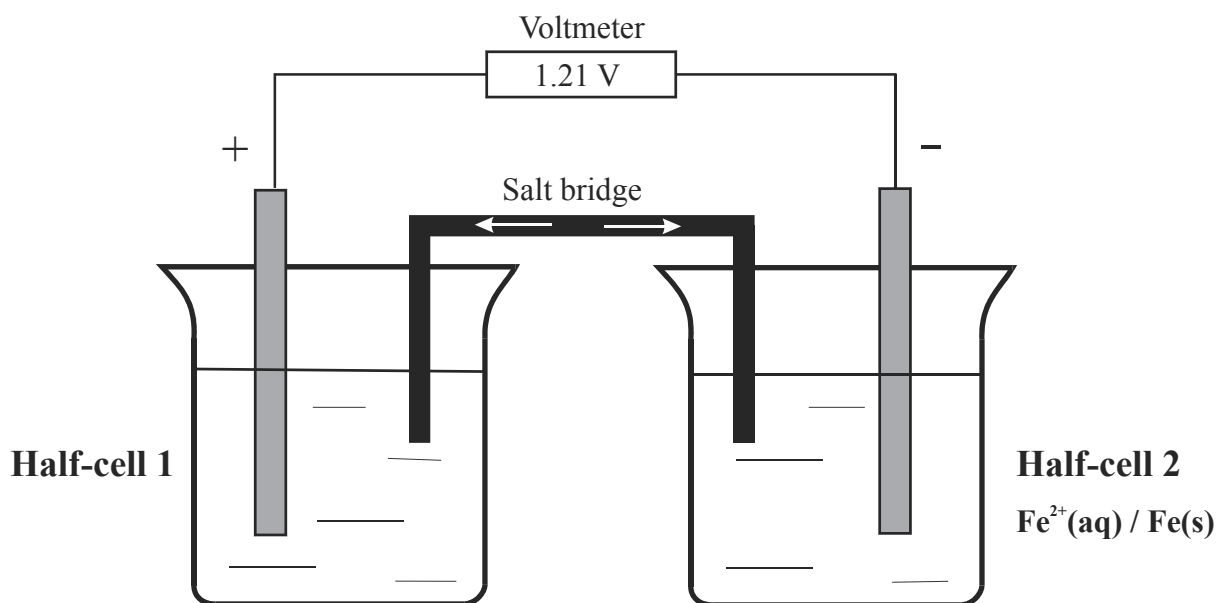


would be

- A. 1635 KJ.
- B. 2135 kJ.
- C. 2490 kJ.
- D. 4270 kJ.

Question 6

Consider the galvanic cell represented below

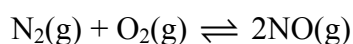


Which of the following sets of species would most likely be present in Half-cell 1?

- A. O₂(g), H⁺(aq), H₂O(l), Pt.
- B. Al³⁺(aq), Al(s).
- C. Fe³⁺(aq), Fe²⁺(aq), Pt.
- D. H⁺(aq), H₂(g), Pt.

Question 7

At high temperature nitrogen and oxygen can react to produce nitrogen(II) oxide, NO. All three species can exist in equilibrium as described by the equation



At 2300 K the equilibrium constant, K_c , for this equilibrium is 1.7×10^{-3} .

In a sealed rigid container at 2300 K, the concentration of N₂ is 0.50 mol L⁻¹, the concentration of O₂ is 0.25 mol L⁻¹ and the concentration of NO is 4.2×10^{-3} mol L⁻¹.

Which of the following statements about this mixture is correct?

- A. The rates of the forward and reverse reactions are equal.
- B. The forward reaction is proceeding slower than the reverse reaction.
- C. The forward reaction is proceeding faster than the reverse reaction.
- D. The pressure in the container is increasing.

Question 8

When a current of 1.40 A was passed through the electric heater of a bomb calorimeter for 54.5 seconds at a potential difference 6.00 volts, the temperature of the calorimeter rose by 0.390°C.

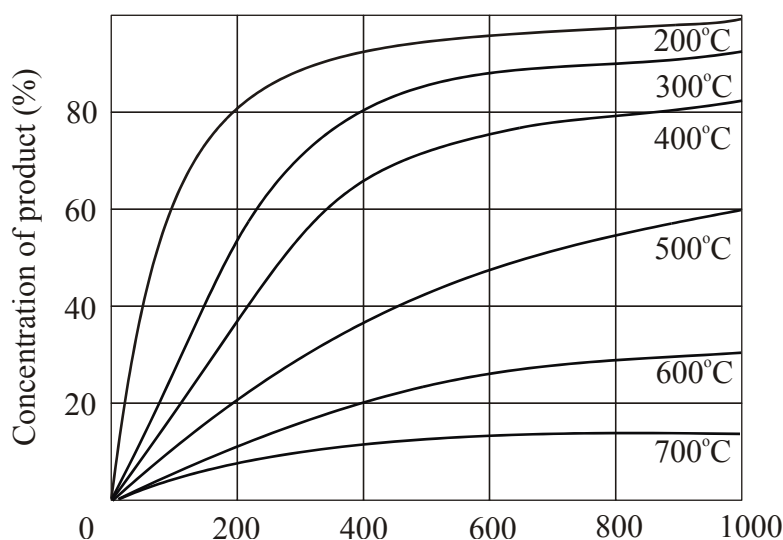
When 9.50×10^{-4} mol of a gaseous hydrocarbon gas underwent combustion in the calorimeter, the temperature rose from 21.45 to 22.17°C.

The hydrocarbon used was

- A. butane.
- B. ethane.
- C. propane.
- D. methane.

Question 9

Temperature and pressure are key factors in the production of common chemicals via equilibrium reactions. The data below show the relationships between percentage yield of product, temperature and pressure for the industrial production of a particular chemical.



According to these data,

- A. the forward reaction is exothermic, and the product is on the side with fewer particles.
- B. the forward reaction is endothermic, and the product is on the side with more particles.
- C. the forward reaction is exothermic, and the product is on the side with more particles.
- D. the forward reaction is endothermic, and the product is on the side with fewer particles.

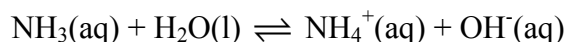
Question 10

The half-equation for the reaction occurring at the negative electrode, during the production of fluorine by electrolysis of a fluoride salt, could be

- A. $2\text{F}^-(\text{aq}) \rightarrow \text{F}_2(\text{g}) + 2\text{e}^-$
- B. $2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$
- C. $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$
- D. $\text{K}^+(\text{l}) + \text{e}^- \rightarrow \text{K}(\text{l})$

Question 11

Ammonia is a weak base, which ionises in water according to the equation

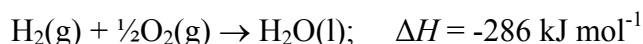


If 20 mL of 0.10 M $\text{NH}_3(\text{aq})$ is added to 50 mL of water, which of the following correctly describes the expected change in pH and number of ammonium ions present as a result of the dilution?

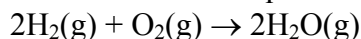
- A. pH decreases, $N(\text{NH}_4^+)$ increases.
- B. pH decreases, $N(\text{NH}_4^+)$ decreases.
- C. pH increases, $N(\text{NH}_4^+)$ increases.
- D. pH increases, $N(\text{NH}_4^+)$ decreases.

Question 12

Hydrogen reacts with oxygen to form water according to the thermochemical equation



Hence ΔH for the equation

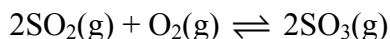


is most likely to be

- A. $+572 \text{ kJ mol}^{-1}$
- B. -572 kJ mol^{-1}
- C. -484 kJ mol^{-1}
- D. $+484 \text{ kJ mol}^{-1}$

Question 13

When $\text{SO}_2(\text{g})$ reacts with $\text{O}_2(\text{g})$ to produce $\text{SO}_3(\text{g})$, equilibrium is established, according to



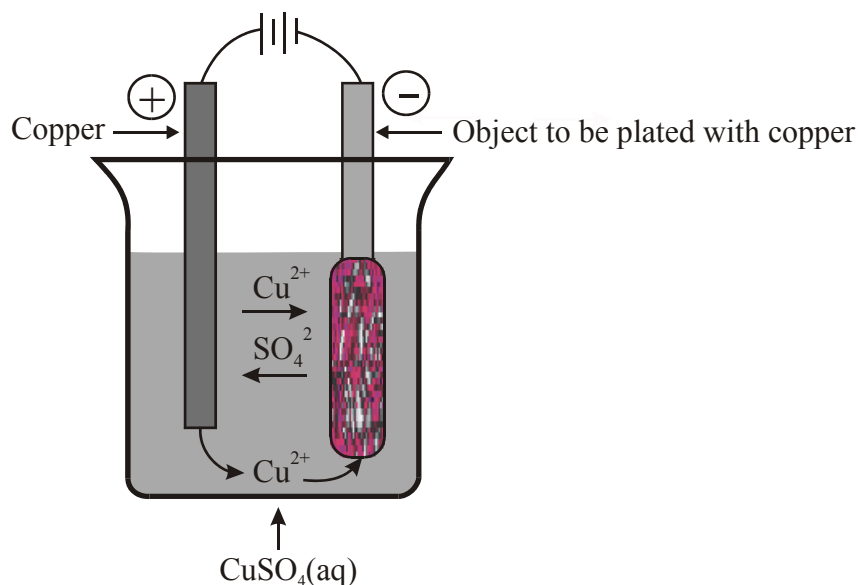
At 1200 K, the equilibrium constant is 0.11 M^{-1}

For the equilibrium $\text{SO}_3(\text{g}) \rightleftharpoons \text{SO}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g})$ at 1200 K, the numerical value of the equilibrium constant will be

- A. 0.33
- B. 3.0
- C. 9.0
- D. 1.2×10^2

Question 14

An electrolytic cell used to electroplate objects with copper is represented below.



At regular intervals the (+) copper electrode is replaced with a new copper electrode. If the (+) copper electrode was mistakenly replaced with a carbon (graphite) electrode,

- A. the pH of the solution would decrease.
- B. bubbles of gas would immediately appear at the (-) electrode.
- C. the concentration of $\text{Cu}^{2+}(\text{aq})$ would not be affected.
- D. the carbon electrode would become coated with copper.

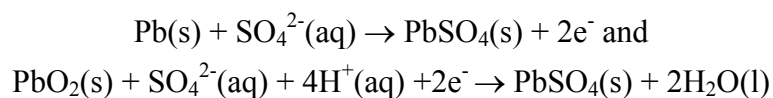
Question 15

If 20 mL of 0.050 M hypochlorous acid is added to 40 mL of water, the pH of resulting solution at 25°C should be closest to

- A. 1.3
- B. 1.8
- C. 4.4
- D. 4.7

Question 16

The common 12 V car battery consists of 6 lead-acid cells, all of which convert chemical energy into electrical energy via the electrode reactions.

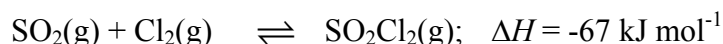


When a car battery is recharged

- A. Pb is produced at the negative electrode of each cell.
- B. the pH in each cell increases.
- C. PbSO_4 is produced at the positive electrode in each cell.
- D. the changes in the oxidation numbers of lead are from 0 to +2 and +4 to +2.

Question 17

Sulfuryl chloride, SO_2Cl_2 , is produced via a reaction between sulfur dioxide and chlorine, according to the equilibrium

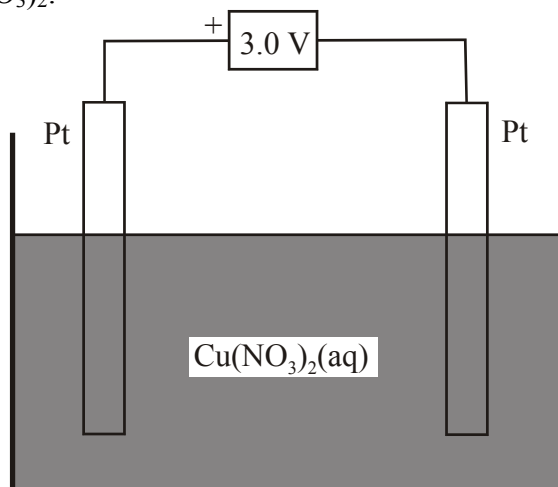


Which of the following changes will increase the yield of thionyl chloride?

- A. Increasing the temperature.
- B. Decreasing the volume of the reaction vessel.
- C. Introducing a new catalyst.
- D. Decreasing the pressure of the equilibrium mixture.

Question 18

In the electrolysis cell shown below, platinum electrodes are placed in 200 mL 0.10 M copper(II) nitrate, $\text{Cu}(\text{NO}_3)_2$.

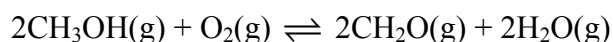


When 9650 C of electric charge is passed through this cell, the maximum amount of gaseous product(s) formed would be

- A. 0.0250 mol O_2 .
- B. 0.0500 mol H_2 .
- C. 0.0250 mol O_2 and 0.030 mol H_2 .
- D. zero, no gaseous products are formed.

Question 19

Methanal, CH_2O , can be produced from methanol, CH_3OH , according to the equilibrium



In an investigation of this equilibrium, a mixture of 3 mol CH_3OH and 3 mol O_2 is allowed to reach equilibrium.

The amount of methanal, CH_2O , present at equilibrium would be

- A. 1.5 mol.
- B. less than 3 mol.
- C. 3 mol.
- D. 6 mol.

Question 20

The large scale generation of electrical energy in coal fired power stations involves the following energy changes:

1. Chemical to thermal
2. Thermal to thermal
3. Thermal to mechanical
4. Mechanical to electrical

Which of the energy changes listed is not associated with the generation of electrical energy in nuclear power stations?

- A. 1
- B. 2
- C. 3
- D. 4

End of Section A

VCE Chemistry 2009 Year 12 Trial Exam Unit 4

Short Answer Questions - Section B

*Section B consists of 6 short answer questions.
You should answer all of these questions.
This section is worth approximately 75 per cent of the total marks available.
The marks allotted are shown at the end of each part of each question.
Questions should be answered in the spaces provided.*

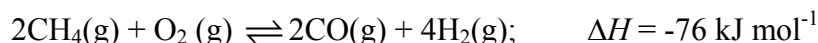
Question 1

Converting natural gas into a mixture of carbon monoxide and hydrogen, known as *synthesis gas* (Syngas), is an important intermediate step in many existing and emerging energy conversion technologies. Syngas can be used to produce methanol, used in direct methanol fuel cells (DMFCs).

- a. Syngas is commonly produced by steam methane reforming (SMR).



Under investigation is the production of Syngas by catalytic partial oxidation (CPOX) of methane



- i. Explain why there may be a significant cost benefit of producing Syngas by CPOX rather than SMR.

(2 marks)

- ii. Explain why a rate/yield conflict may exist in the CPOX process.

(2 marks)

- iii. Identify one applied condition which could, according to Le Chatelier's principle, increase the yield of Syngas in both SMR and CPOX.

(1 mark)

- iv. Give the name and chemical formula of the ‘industrial chemical’ you studied in detail during Unit 4, and write a balanced equation for one of the equilibria associated with its production.

(2 marks)

- v. Identify one applied condition which could theoretically increase the yield in the equilibrium described in (iv) and the production of Syngas by CPOX.

(1 mark)

- vi. Methanol is produced by passing Syngas over a Cu/ZnO/Al₂O₃ catalyst at 250°C. Write a balanced equation for the production of methanol from Syngas.

(1 mark)

- b. The technology behind Direct Methanol Fuel Cells (DMFC) is still in the early stages of development, but it has been successfully demonstrated powering mobile phones and laptop computers—potential target end uses in future years. The electrolyte in a DMFC is a polymer through which the charge carrier, H⁺, can move. The end products of the reaction in a DMFC are the same for the combustion of methanol.

- i. Write a balanced half-equation for the reaction occurring at the (-) electrode in a DMFC.

(1 mark)

- ii. Write a balanced half-equation for the reaction occurring at the (+) electrode in a DMFC.

(1 mark)

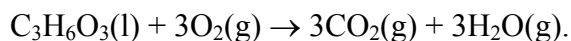
- iii. In terms of the fuel used, suggest one advantage that DMFCs may have over a H₂/O₂ fuel cell.

(1 mark)

Total 12 marks

Question 2

A bomb calorimeter was used to determine the ΔH for the combustion of lactic acid, $\text{C}_3\text{H}_6\text{O}_3$. The equation for the reaction is



The following procedure was followed:

1. The calorimeter was calibrated by allowing a 5.000 g sample of ethanol to react completely with oxygen in the calorimeter. The energy released caused the temperature of the calorimeter and its contents to change from 23.11°C to 53.19°C .
 2. A 5.865 sample of pure lactic acid was then reacted completely with oxygen in the calorimeter. The energy released caused the temperature of the calorimeter and its contents to change from 23.1°C to 41.6°C .
- a. Calculate the calibration factor.

(3 marks)

- b. Calculate the energy released from the combustion of 5.865 g of lactic acid.

(2 marks)

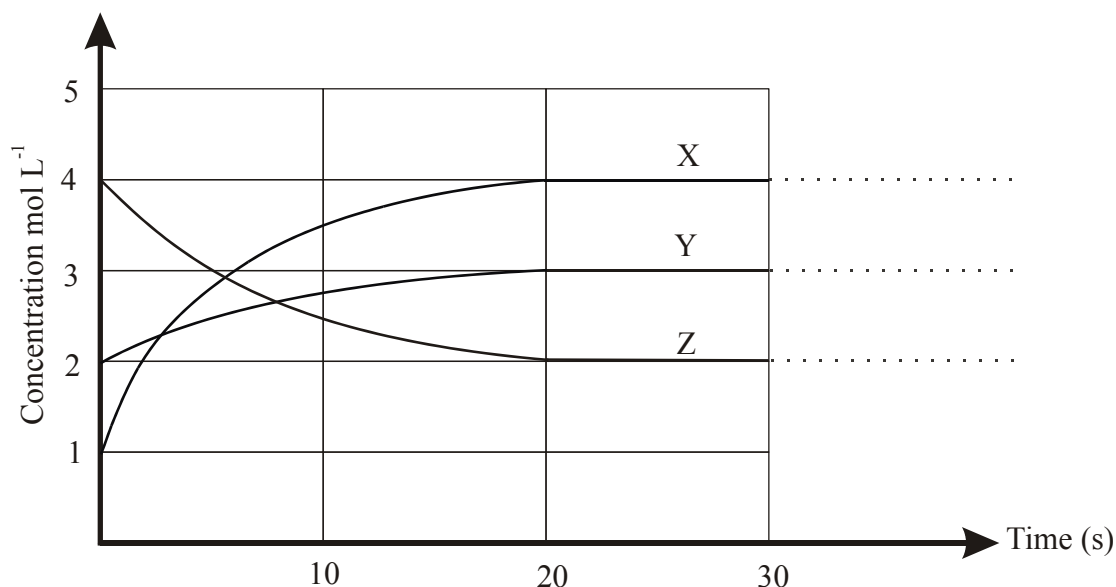
- c. Determine ΔH for the reaction $\text{C}_3\text{H}_6\text{O}_3(\text{aq}) + 3\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$.

(3 marks)

Total 8 marks

Question 3

The graph below shows the concentrations, as a function of time, for a mixture of gases, X, Y and Z, added to a one litre container and allowed to reach equilibrium.



- a. How long does the mixture take to reach equilibrium?

(1 mark)

- b. i. Use the information provided on the graph to complete the table below.

	Initial Concentration	Equilibrium Concentration	How the concentration changes as the system moves to equilibrium
X			
Y			
Z			

(3 marks)

- ii. Use the information in the table in (i) to deduce the equation for the equilibrium reaction.

(2 marks)

- c. Write the equilibrium law expression and calculate the value of the equilibrium constant.

(2 marks)

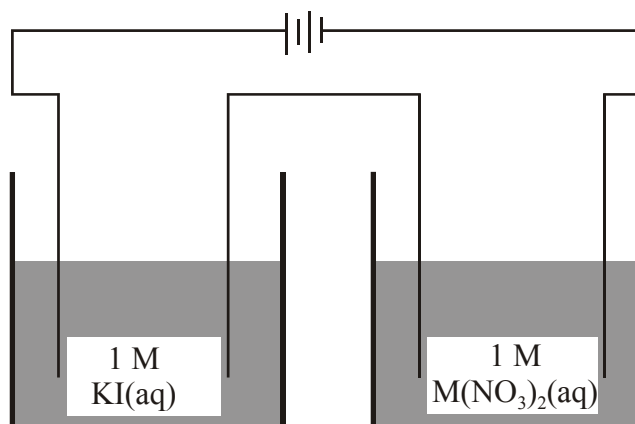
- d. At 30 seconds, the volume of the container was doubled. On the concentration time graph, sketch how the concentration of X would change during and after this volume increase.

(2 marks)

Total 10 marks

Question 4

A current of 2.5 A is passed for 50 minutes through two cells connected in series. As indicated on the diagram below, these cells contained 1 M solutions of potassium iodide and a metal nitrate, $M(NO_3)_2$, respectively. Each cell also contains a pair of *platinum electrodes*.



- a. Calculate the charge, in Coulombs, passed through each cell.

(1 mark)

- b. Calculate the amount of electrons, in mole, passed through each cell.

(1 mark)

- c. During the electrolysis, a metal is deposited on one of the electrodes in the cell containing the metal nitrate, but not in the cell containing potassium iodide. What does this suggest about the relative positions on the electrochemical series of potassium metal and the metal in the metal nitrate?

(2 marks)

- d. Calculate the mass of the product formed from the reaction at the positive electrode in the potassium iodide cell.

(2 marks)

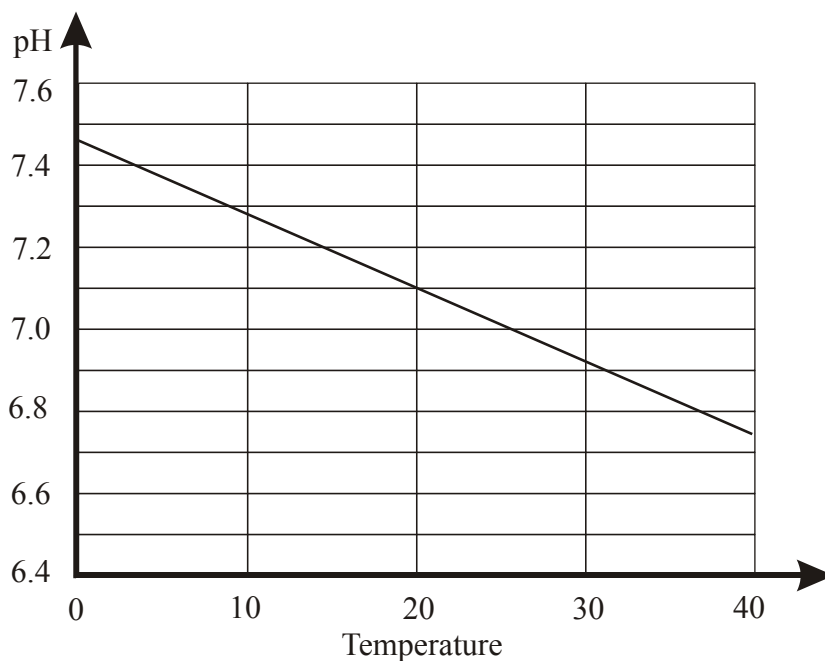
- e. During the electrolysis, 2.48 g of the metal is deposited on the cathode in the metal nitrate cell. Identify the metal.

(2 marks)

Total 8 marks

Question 5

- a. The graph below shows the variation in pH of pure water with temperature.



- i. Explain why the pH of pure water changes with temperature as shown in this graph.

(2 marks)

- ii. Calculate the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in pure water at 35°C.

(2 marks)

- b. Calculate the pH of 0.10 M propanoic acid.

(3 marks)

- c. i. Show that when 10 mL of 0.10 M HCl(aq) is diluted to 100 mL with water the pH of the solution increases by 1.

(2 marks)

- ii. Show that when 10 mL of 0.10 M propanoic acid is diluted to 100 mL with water, the pH increases by less than 1.

(3 marks)

- iii. Explain why the change in pH is less when 10 mL of 0.10 M propanoic acid is diluted to 100 mL with water, than when 10 mL of 0.10 M hydrochloric acid is diluted to 100 mL with water.

(3 marks)

Total 15 marks

Question 6

Standard electrode potentials associated with half-equations on the electrochemical series, are determined from the cell voltage generated when the half-cell is connected to the standard hydrogen half-cell, under standard conditions, in a functioning galvanic cell.

- a. Describe the components and conditions present in a standard hydrogen half-cell.

(2 marks)

- b. Write an overall equation describing the cell reaction occurring when a standard hydrogen half-cell is connected to a $\text{Cl}_2(\text{g})/\text{Cl}^-(\text{aq})$ half-cell in an operating galvanic cell.

(1 mark)

- c. In a different galvanic cell containing the standard hydrogen half-cell, the pH in the standard hydrogen half-cell decreases as the cell delivers energy. Explain how this knowledge enables you to deduce the sign of the electrode in the standard hydrogen half-cell.

(2 marks)

- d. In the alternative version of the electrochemical series, standard electrode potentials are determined by using the $\text{Cu}^{2+}(\text{aq})/\text{Cu}(\text{s})$ half-cell as the reference half-cell. Assuming the same standard conditions apply, what would be standard electrode potential of the $\text{Ni}^{2+}(\text{aq})/\text{Ni}$ half-cell in this alternative electrochemical series? Explain your reasoning.

(2 marks)

- e. Referring to standard electrode potentials, explain why aluminium cannot be produced by the electrolysis of an aqueous solution of Al^{3+} ions, i.e. $\text{Al}^{3+}(\text{aq})$

(2 marks)

Total 9 marks

End of Section B

End of Trial Exam

Suggested Answers

VCE Chemistry 2009 Year 12 Trial Exam Unit 4

Multiple Choice Answers - Section A

Question 1

- C. The graph shows that the quantity on the vertical axis increases with time, but the rate of increase – as represented by the gradient of the curve – decreases with time. Since HCl(aq) is a strong acid, the reaction may be represented by the ionic equation $\text{Zn(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g}); \Delta H = -154 \text{ kJ mol}^{-1}$. Consider the alternatives:
- A. As the reaction proceeds, the number of collisions per second between the reactant particles will decrease as the $[\text{H}^+]$ decreases.
 - B. The rate of reaction decreases with time as the $[\text{H}^+]$ decreases.
 - C. Since the reaction is **exothermic** the **temperature of the solution will increase**, rapidly initially and then more slowly as the rate of reaction decreases.
 - D. Number of ions decreases with time as H^+ ions are converted into H_2 molecules.

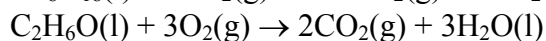
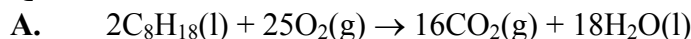
Question 2

- D. According to oxidation numbers, **Zn is oxidised** (from 0 in Zn to +2 in Zn(OH)_2) and **O is reduced** (from 0 in O_2 to -2 in Zn(OH)_2). Since oxidation occurs at the negative electrode when a galvanic cell is converting chemical energy to electrical energy, then **Zn must be the negative electrode**. The half-equations can be deduced from the electrochemical series
- Anode (-)** $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ oxidation
- Cathode (+)** $\text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$ reduction
- Multiply the oxidation half-equation by 2 and add the reduction half-equation to get the overall equation $2\text{Zn(s)} + \text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Zn(OH)}_2(\text{s})$

Question 3

- D. The rate of a chemical reaction is determined by the **proportion of reactant molecule collisions that have energy greater than the activation energy for the reaction**. This proportion is reflected, on a kinetic energy distribution curve, as the **area under the curve to the right of the activation energy**. This area is **greatest for reaction 4**, hence this reaction should occur at the fastest rate.

Question 4



Energy released = 2000 kJ

According to the Data Book

$\Delta H_c(\text{C}_8\text{H}_{18}) = -5464 \text{ kJ mol}^{-1}$; $\Delta H_c(\text{C}_2\text{H}_6\text{O}) = -1364 \text{ kJ mol}^{-1}$

$n(\text{C}_8\text{H}_{18}) \text{ reacting} = 2000 \text{ kJ} / 5464 \text{ kJ mol}^{-1}$
 $= 0.3660 \text{ mol}$

$n(\text{CO}_2) \text{ produced from C}_8\text{H}_{18} = (16/2) \times n(\text{C}_8\text{H}_{18})$
 $= 8 \times 0.3660$
 $= \mathbf{2.928 \text{ mol}}$

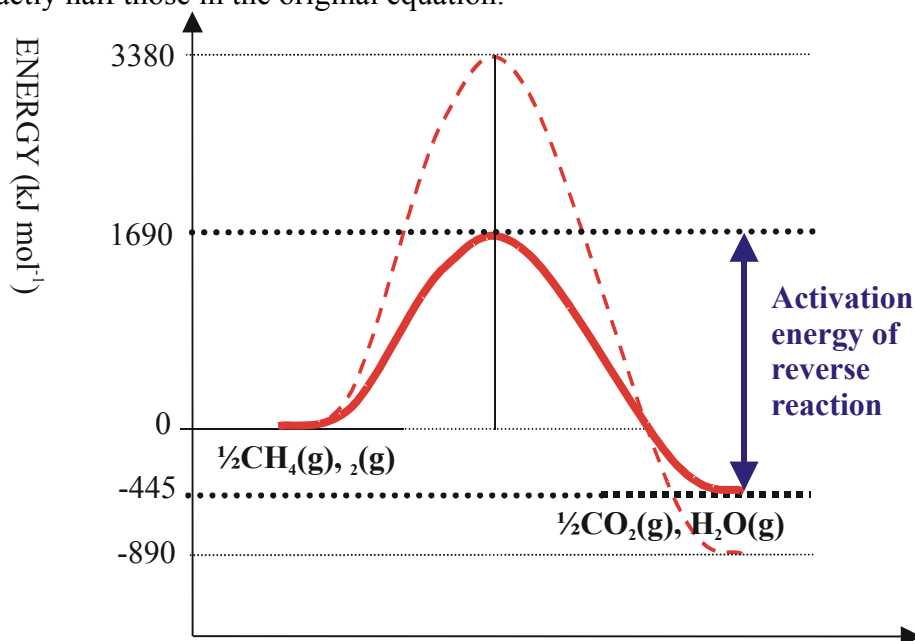
$n(\text{C}_2\text{H}_6\text{O}) \text{ reacting} = 2000 \text{ kJ} / 1364 \text{ kJ mol}^{-1}$
 $= 1.466 \text{ mol}$

$n(\text{CO}_2) \text{ produced from C}_2\text{H}_6\text{O} = 2 \times n(\text{C}_2\text{H}_6\text{O})$
 $= 2 \times 1.466$
 $= \mathbf{2.933 \text{ mol}}$

Ratio $n(\text{CO}_2) \text{ produced from C}_8\text{H}_{18} : n(\text{CO}_2) \text{ produced from C}_2\text{H}_6\text{O} = 2.928 : 2.933$
 $= 2.928/2.933 : 1$
 $= \mathbf{1 : 1}$

Question 5

- B. The energy profile for $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ can be adjusted to give the energy profile for $\frac{1}{2}\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \frac{1}{2}\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$, by halving all the values on the original energy profile, and the ΔH value, since all the mole ratios are exactly half those in the original equation.



From the energy profile for $\frac{1}{2}\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \frac{1}{2}\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$, the activation energy for $\frac{1}{2}\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \frac{1}{2}\text{CH}_4(\text{g}) + \text{O}_2(\text{g})$, i.e. the reverse reaction is $1690 - (-445) = \mathbf{2135 \text{ kJ}}$

Alternatively:

From the energy profile for $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
the activation energy for the reaction $\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \rightarrow \text{CH}_4(\text{g}) + 2\text{O}_2(\text{g})$ is
 $3380 - (-890) = 4270 \text{ kJ mol}^{-1}$.

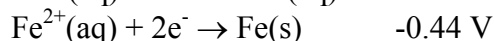
Hence the activation energy for the reaction $\frac{1}{2}\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \frac{1}{2}\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
is $4270 / 2 = \mathbf{2135 \text{ kJ mol}^{-1}}$.

Question 6

- C. In a galvanic cell electrons move spontaneously from the anode (-) to the cathode (+).
So the electrode signs indicate that oxidation is occurring in the Fe^{2+}/Fe half-cell i.e.
 $\text{Fe}(\text{s}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{e}^-$.

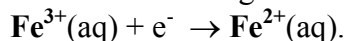
Since electrons leave the site of oxidation and go to the site of reduction, the left-hand half-cell must contain an oxidant which is 1.21 V higher up on the electrochemical series than the reductant $\text{Fe}(\text{s})$.

Scanning the electrochemical series and looking for a difference of 1.21 V indicates that the oxidant must be $\text{Fe}^{3+}(\text{aq})$, according to



$$E(\text{cell}) = 0.77 - (-0.44) \\ = 1.21 \text{ V}$$

Reaction occurring in the half-cell 1 is



Since all the species in the half-equation are ions, a platinum electrode is used in the half-cell.

Alternatively:

Since $E(\text{cell}) = E^\circ(\text{oxidant half-cell}) - E^\circ(\text{reductant half-cell})$

$$1.21 \text{ V} = E^\circ(\text{oxidant}) - E^\circ(\text{Fe}^{2+}(\text{aq})/\text{Fe}(\text{s}))$$

$$= E^\circ(\text{oxidant}) - (-0.44 \text{ V})$$

$$= E^\circ(\text{oxidant}) + 0.44 \text{ V}$$

$$\text{Hence } E^\circ(\text{oxidant}) = 1.21 - 0.44 \\ = 0.77 \text{ V}$$

From the electrochemical series $\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq}); E^\circ = +0.77 \text{ V}$

Question 7

- C. To decide whether the system is or is not at equilibrium, it is necessary to calculate the value of the reaction quotient (Q) and compare it with the equilibrium constant.

$$Q = [\text{NO}]^2 / \{[\text{N}_2] \times [\text{O}_2]\} \\ = (4.2 \times 10^{-3})^2 / \{0.50 \times 0.25\} \\ = 1.4 \times 10^{-4}$$

Q (1.4×10^{-4}) is less than K_c (1.7×10^{-3}) so, for the system to get to equilibrium, Q must increase until its value is the same as the equilibrium constant.

Since Q is effectively the ratio of product concentrations to reactant concentrations, the reaction must proceed in the forward direction to give more products and less reactants. In order for the reaction to move to the right, the **rate of the forward reaction must be greater than the rate of the reverse reaction.**

Irrespective of the reaction direction, or if the system is at equilibrium, the pressure in the container will remain the same because the total number of particles present does not change.

Question 8

D. Calibration Factor = Energy added during heating / ΔT_c
 $= VIt / \Delta T_c$
 $= 6.00 \times 1.40 \times 54.5 / 0.390$
 $= 457.8 / 0.390$
 $= 1174 \text{ J } ^\circ\text{C}^{-1}$

Energy released by hydrocarbon = $CF \times \Delta T_r$
 $= 1174 \times (22.17 - 21.45)$
 $= 1174 \text{ J } ^\circ\text{C}^{-1} \times 0.72 ^\circ\text{C}$
 $= 845 \text{ J}$

Energy per mol hydrocarbon = $845 \text{ J} / 9.50 \times 10^{-4} \text{ mol}$
 $= 8.90 \times 10^5 \text{ J}$
 $= 8.90 \times 10^2 \text{ kJ}$
 $= \mathbf{890 \text{ kJ}}$

So the hydrocarbon has a molar enthalpy (ΔH_c) of combustion of $\mathbf{-890 \text{ kJ mol}^{-1}}$.
According to Table 13 in the Data Book. $\Delta H_c(\text{methane}) = \mathbf{-889 \text{ kJ mol}^{-1}}$

Question 9

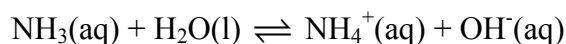
- A. The supplied data show clearly that the ‘**yield of product**’ **increases**
– as the **temperature decreases**
– as the **pressure increases**
If the yield is increasing as the temperature is decreasing, the **forward reaction** is favoured by low temperature, i.e. is **exothermic**.
If the yield is increasing as the pressure increases, then since the equilibrium system responds to a pressure increase, by striving to counteract the pressure increase by moving to decrease the pressure, the position of equilibrium is pushed to the **side with fewer particles**, so the **product** must be on that side.

Question 10

- D. Fluorine is produced by the electrolysis of salts containing the fluoride ions. However according to the electrochemical series, particularly
 $\text{F}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{F}^-(\text{aq}) \quad +2.87 \text{ V}$
 $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l}) \quad +1.23 \text{ V}$
aqueous solutions of fluoride salts cannot be used because **$\text{H}_2\text{O}(\text{l})$ is a stronger reductant than $\text{F}^-(\text{aq})$ and would be preferentially oxidised.**
Fluorine is produced by the electrolysis of molten salts such as $\text{KF}(\text{l})$, for which the electrode half-equations are
(+) oxidation: $2\text{F}^-(\text{l}) \rightarrow \text{F}_2(\text{g}) + 2\text{e}^-$
(-) reduction: $\text{K}^+(\text{l}) + \text{e}^- \rightarrow \text{K}(\text{l})$

Question 11

- A. The addition of the 20 mL of 0.10 M $\text{NH}_3(\text{aq})$ to 50 mL of water pushes the equilibrium



to the right.

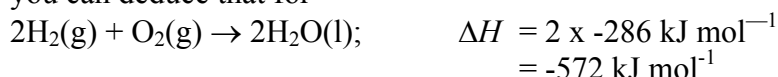
This is consistent with Le Chatelier's principle – the system moves to compensate for the extra reactant, H_2O , by favouring the forward reaction until equilibrium is again reached.

So the **number of NH_4^+ ions increases** and the number of OH^- ions (the link to pH) also increases.

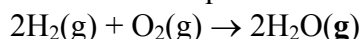
However since the total volume has increased, the **increase in $n(\text{OH}^-)$ will not be enough to compensate for the volume increase**. The $[\text{OH}^-]$ will be **lower** at the new equilibrium and consequently the **pH will be lower**.

Question 12

- C. Given $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta H = -286 \text{ kJ mol}^{-1}$
you can deduce that for



However the equation for which the ΔH is required is

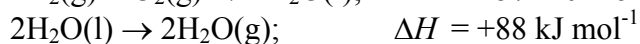
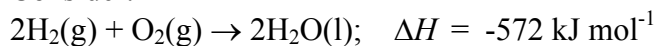


The **formation of $\text{H}_2\text{O}(\text{g})$ from $\text{H}_2(\text{g})$ and $\text{O}_2(\text{g})$ releases less energy per mole than the formation of $\text{H}_2\text{O}(\text{l})$** . This is consistent with the fact that we add energy to boil water, i.e. the reaction $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$ is endothermic, and hence $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$ is exothermic.

So for $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}); \Delta H = -484 \text{ kJ mol}^{-1}$ is more accurate.

It is also consistent with $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta H = -44 \text{ kJ mol}^{-1}$

Consider:

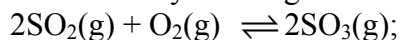


Hence

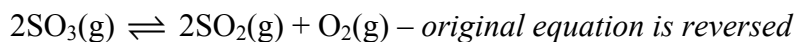


Question 13

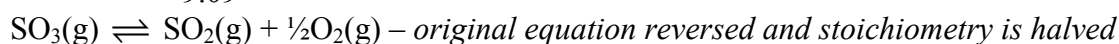
- B. Consider the variations that occur in the equilibrium law expression as the equation stoichiometry is changed.



$$K_1 = [\text{SO}_3]^2 / \{[\text{SO}_2]^2[\text{O}_2]\} \\ = 0.11$$



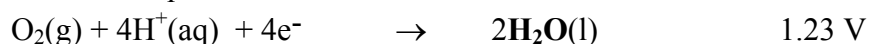
$$K_2 = \{[\text{SO}_2]^2[\text{O}_2]\} / [\text{SO}_3]^2 \\ = 1 / ([\text{SO}_3]^2 / \{[\text{SO}_2]^2[\text{O}_2]\}) \\ = 1 / K_1 \\ = 1 / 0.11 \\ = 9.09$$



$$K_3 = \{[\text{SO}_2][\text{O}_2]^{1/2}\} / [\text{SO}_3] \\ = (\{[\text{SO}_2]^2[\text{O}_2]\} / [\text{SO}_3]^2)^{1/2} \\ = K_2^{1/2} \\ = \sqrt{K_2} \\ = \sqrt{9.09} \\ = 3.0$$

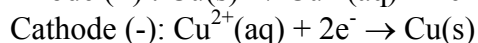
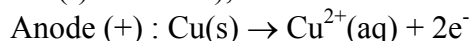
Question 14

- A. The relevant half-equations from the electrochemical series are



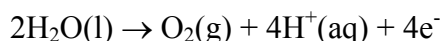
Because electroplating is electrolysis, oxidation occurs at the (+) electrode and reduction occurs at the (-) electrode.

On the basis of the reaction between the strongest oxidant and strongest reductant principle (and the clue of $\text{Cu}^{2+}(\text{aq})$ ions released at the (+) electrode and consumed at the (-) electrode), the reactions occurring in the normal copper-plating cell are



If the copper anode, i.e. the (+) electrode, is replaced with a carbon electrode, the only reductant present at the carbon anode will be H_2O .

So instead of Cu being oxidised at the (+) electrode, H_2O will be oxidised according to



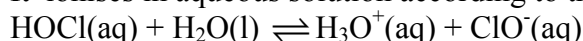
Thus bubbles of oxygen gas will be produced at the (+) electrode. Also, because of the increase in $[\text{H}^+]$ in the solution, the **pH** will **decrease**.

$\text{Cu}(\text{s})$ would still be produced at the (-) electrode until the $c(\text{Cu}^{2+})$ has dropped to an unworkable level.

Question 15

- D. Hypochlorous acid, HOCl(aq), is listed in Table 12 of the Data Book as a weak acid with $K_a = 2.9 \times 10^{-8}$.

It ionises in aqueous solution according to the equilibrium



For which $K_a = [\text{H}_3\text{O}^+]_e [\text{ClO}^-]_e / [\text{HOCl}]_e$

Since 20 mL of 0.050 M HOCl(aq) is added to 40 mL of water, we need to determine the initial concentration of the diluted solution

$$\begin{aligned} n(\text{HOCl}) \text{ in } 20 \text{ mL } 0.050 \text{ M} &= cV \\ &= 0.050 \times 20 \times 10^{-3} \\ &= 1.0 \times 10^{-3} \text{ mol} \\ c(\text{HOCl}) \text{ in } 60 \text{ mL solution} &= n(\text{HOCl}) / V \\ &= 1.0 \times 10^{-3} / 60 \times 10^{-3} \\ &= 0.0167 \text{ mol L}^{-1} \end{aligned}$$

Assumptions for a weak acid

1. Limited ionisation of HOCl, so that $[\text{HOCl}]_e = [\text{HOCl}]_{\text{initially}} = 0.0167 \text{ M}$

2. $[\text{H}_3\text{O}^+]_e = [\text{ClO}^-]_e$

Since $K_a = 2.9 \times 10^{-8}$

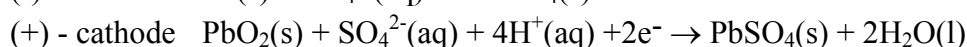
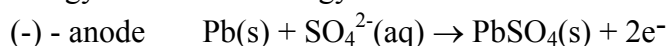
$$\begin{aligned} 2.9 \times 10^{-8} &= [\text{H}_3\text{O}^+]_e [\text{ClO}^-]_e / [\text{HOCl}]_e \\ &= [\text{H}_3\text{O}^+]^2 / 0.0167 \end{aligned}$$

$$\begin{aligned} 0.0167 \times 2.9 \times 10^{-8} &= [\text{H}_3\text{O}^+]^2 \\ [\text{H}_3\text{O}^+]^2 &= 4.84 \times 10^{-10} \\ [\text{H}_3\text{O}^+] &= \sqrt{4.84 \times 10^{-10}} \\ &= 2.20 \times 10^{-5} \end{aligned}$$

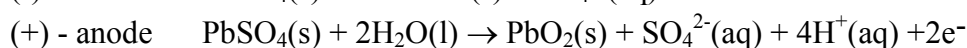
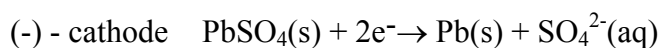
$$\begin{aligned} \text{pH} &= -\log_{10} [\text{H}_3\text{O}^+] \\ &= -\log_{10} (2.20 \times 10^{-5}) \\ &= -(-4.7) \\ &= 4.7 \end{aligned}$$

Question 16

- A. The electrode signs for the discharging reactions in each cell which convert chemical energy to electrical energy are



So the reactions occurring at these electrodes during the recharging reactions in each cell are



Checking the various alternatives

- A. **Pb is produced at the (-) electrode**

- B. pH will decrease as H^+ ions are produced at (+) electrode

- C. PbSO_4 is consumed

- D. Oxidation numbers of Pb change from +2 to +4 at the anode, and +2 to 0 at the cathode

Question 17

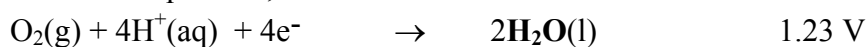
B. Changes that will increase the yield of SO_2Cl_2 by pushing the position of equilibrium to the right and are consistent with Le Chatelier's principle include:

- **decreasing the temperature**, since the forward reaction is exothermic
- **increasing the pressure / decreasing the volume**, since the system moves to produce fewer particles in the larger volume
- **increasing reactant concentration by adding a reactant**
- **decreasing product concentration by removing a product**

However whilst a **catalyst** speeds up the rate of both the forward and reverse reactions, by lowering their respective activation energies, it has **no effect on the position of equilibrium**, hence no effect on the yield of SO_2Cl_2 .

Question 18

C The relevant half-equations, from the electrochemical series are



$$\begin{aligned} n(\text{Cu}^{2+}) \text{ in cell before electrolysis} &= c \times V \\ &= 0.10 \times 200 \times 10^{-3} \\ &= 0.020 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{e}^-) \text{ passed through the cell} &= Q / F \\ &= 9650 / 96500 \\ &= 0.1000 \text{ mol} \end{aligned}$$

Strongest oxidant present in the cell is $\text{Cu}^{2+}(\text{aq})$, and the strongest reductant present is $\text{H}_2\text{O}(\text{l})$. So the reactions at the electrodes would be

Anode (+) reaction: $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$.

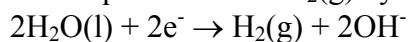
$$\begin{aligned} n(\text{O}_2) \text{ produced} &= n(\text{e}^-) / 4 \\ &= \mathbf{0.025 \text{ mol}} \end{aligned}$$

Cathode (-) reaction: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$

$$\begin{aligned} n(\text{Cu}) \text{ produced} &= n(\text{Cu}^{2+}) \\ &= 0.020 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{e}^-) \text{ used to produce Cu} &= 2 \times n(\text{Cu}) \text{ produced} \\ &= 0.040 \text{ mol} \end{aligned}$$

However this leaves $0.1000 - 0.040 = 0.060 \text{ mol e}^-$ for another cathode reaction, i.e. for the production of $\text{H}_2(\text{g})$ by the reduction of H_2O .



$$\begin{aligned} n(\text{H}_2) \text{ produced} &= \frac{1}{2} \times n(\text{e}^-) \text{ used} \\ &= \frac{1}{2} \times 0.060 \\ &= \mathbf{0.030 \text{ mol}} \end{aligned}$$

So the maximum amounts of gaseous products that could be formed would be **0.025 mol O_2** (at the anode) **and 0.030 mol H_2** (at the cathode).

Question 19

- B.** If the reaction $2\text{CH}_3\text{OH}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CH}_2\text{O}(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ proceeded to completion, 2 mol CH_3OH reacts with 1 mol O_2 to produce 2 mol CH_2O .
Hence if 3 mol CH_3OH was mixed with 3 mol O_2 and the reaction proceeded to completion, 3 mol CH_3OH would react with 1.5 mol O_2 to produce 3 mol CH_2O .
However this would result in no CH_3OH remaining, which cannot be the case if the system reaches equilibrium. Hence when equilibrium has been reached, less than 3 mol CH_3OH must have been converted to CH_2O , **and less than 3 mol CH_2O would be present at equilibrium.**
Although 1.5 mol is less than 3 mol, alternative A is not correct because there was not enough information supplied to determine exactly how much CH_2O is present at equilibrium.

Question 20

- A.** The fundamental difference in a nuclear power station is that the first energy change is the **conversion of nuclear energy (in uranium nuclei) to thermal energy.**
The energy change happens as **large uranium nuclei are split into smaller nuclei of different elements.**
In the conversion of **chemical energy to thermal energy**, the atoms are rearranged as **chemical bonds are broken and formed, but there is no change in the elements present.**
In a nuclear power station, the energy changes are
1. nuclear to thermal – in the **reactor**
 2. thermal to thermal – **production of steam** by boiling water
 3. thermal to mechanical – steam driving the **turbine**
 4. mechanical to electrical – in the **generator**

Short Answer (Answers) - Section B

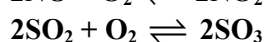
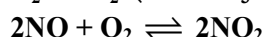
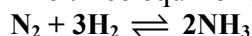
Question 1

- a. i. The CPOX reaction is **exothermic** ($\Delta H < 0$), whereas the SMR process is **endothermic** ($\Delta H > 0$). ❶
Hence the **heating costs should be less with the CPOX process**. Also the **heat released in the CPOX process can be used elsewhere in the process**, e.g. to heat incoming gases. ❶
- ii. Because the CPOX equilibrium is exothermic, the **forward reaction** and yield of Syngas is **favoured by low temperatures**. ❶ However since reaction rates are slower at lower temperatures then **the rate of production of Syngas may be too slow**. ❶
- iii. **Lower pressure**. ❶ Since both processes have more particles on the product side of the equilibrium, the position of equilibrium will shift to the right as the system adjusts to a pressure decrease.

iv. v.

❶ Ammonia – NH_3	Sulfuric Acid – H_2SO_4	Ethene – C_2H_4	Nitric Acid – HNO_3
❶ $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$	$2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$	$\text{C}_3\text{H}_8 \rightleftharpoons \text{C}_2\text{H}_4 + \text{CH}_4$	$2\text{NO} + \text{O}_2 \rightleftharpoons 2\text{NO}_2$
❶ Lower temperature	Lower temperature	Decrease pressure	Lower temperature

The three equilibria



like the CPOX equilibrium all have $\Delta H < 0$ and so the forward reaction is favoured by low temperature. However unlike the CPOX equilibrium, the forward reaction is favoured by high pressure.

The equilibrium $\text{C}_3\text{H}_8 \rightleftharpoons \text{C}_2\text{H}_4 + \text{CH}_4$ has $\Delta H > 0$ and so the forward reaction is favoured by high temperature. However, like the CPOX equilibrium it has more particles on the product side and so the forward reaction is favoured by low pressure.

- vi. $\text{CO(g)} + 2\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_3\text{OH(g)}$ ❶
- b. i. $\text{CH}_3\text{OH(l)} + \text{H}_2\text{O(l)} \rightarrow \text{CO}_2\text{(g)} + 6\text{H}^+\text{(aq)} + 6\text{e}^-$ ❶
- ii. $\text{O}_2\text{(g)} + 2\text{H}^+\text{(aq)} + 2\text{e}^- \rightarrow 2\text{H}_2\text{O(l)}$ ❶
- iii. **Liquid methanol is easier to transport and store** than gaseous hydrogen. ❶
There is also a significant issue with the ‘explosiveness’ of hydrogen

Question 2

- a. From Data Book; $\Delta H_c(\text{ethanol}) = -1364 \text{ kJ mol}^{-1}$
 $n(\text{C}_2\text{H}_5\text{OH})$ used in calibration $= 5.000 \text{ g} / 46.0 \text{ g mol}^{-1}$
 $= 0.109 \text{ mol}$
 Energy released $= 0.109 \text{ mol} \times 1364 \text{ kJ mol}^{-1}$
 $= 148 \text{ kJ}$ ❶
 Temperature change $= 53.19 - 23.11$
 $= 30.08^\circ\text{C}$ ❶
 Calibration factor $= 148 \text{ kJ} / 30.08^\circ\text{C}$
 $= 4.93 \text{ kJ } ^\circ\text{C}^{-1}$ ❶

- b. Energy released = calorimeter constant x temperature change
 $= 4.93 \text{ kJ } ^\circ\text{C}^{-1} \times (41.6 - 23.1)^\circ\text{C}$ ❶
 $= 4.93 \times 18.5 \text{ kJ}$
 $= 91.2 \text{ kJ}$ ❶
- c. $5.865 \text{ g C}_3\text{H}_6\text{O}_3 \rightarrow 91.2 \text{ kJ}$
 To get ΔH you need to determine the energy released per mol $\text{C}_3\text{H}_6\text{O}_3$
 $n(\text{C}_3\text{H}_6\text{O}_3) = m(\text{C}_3\text{H}_6\text{O}_3) / M(\text{C}_3\text{H}_6\text{O}_3)$
 $= 5.865 \text{ g} / (3 \times 12.0 + 6 \times 1.00 + 3 \times 16.0) \text{ g mol}^{-1}$
 $= 5.865 / 90.0 \text{ mol}$
 $= 0.0652 \text{ mol}$ ❶
 Energy per mol $= 91.2 \text{ kJ} / 0.0652 \text{ mol}$
 $= 1.40 \times 10^3 \text{ kJ mol}^{-1}$ ❶
 So for $\text{C}_3\text{H}_6\text{O}_3(\text{aq}) + 3\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$, $\Delta H = -1.40 \times 10^3 \text{ kJ mol}^{-1}$ ❶

Question 3

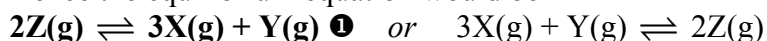
- a. **20 minutes** ❶ - the concentrations stop changing at 20 minutes and the graphs are horizontal.

- b. i.

	Initial Concentration	Equilibrium Concentration	How the concentration changes as the system moves to equilibrium
X	1 mol L ⁻¹	4 mol L ⁻¹	Increases by 3 mol L⁻¹ ❶
Y	2 mol L ⁻¹	3 mol L ⁻¹	Increases by 1 mol L⁻¹ ❶
Z	4 mol L ⁻¹	2 mol L ⁻¹	Decreases by 2 mol L⁻¹ ❶

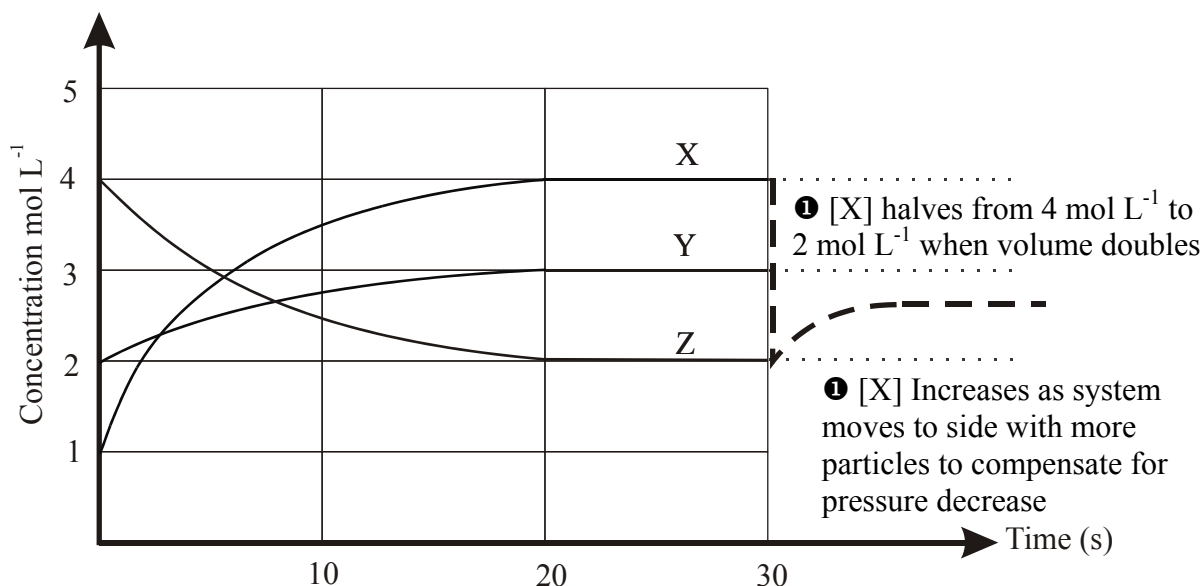
- ii. The concentrations of X and Y increase so more of both is being made as the system moves to equilibrium and Z is being consumed.
 Since it is a 1 L vessel, **2 mol Z has reacted to produce 3 mol of X and 1 mol of Y.** ❶

Hence the equilibrium equation would be



- c. $K = [\text{X}]^3[\text{Y}] / [\text{Z}]^2$ ❶ or $K = [\text{Z}]^2 / [\text{X}]^3[\text{Y}]$
 $= 4^3 \times 3 / 2^2 = 2^2 / (4^3 \times 3)$
 $= 5 \times 10^1 (\text{M}^2)$ ❶ $= 2 \times 10^{-1} (\text{M}^{-2})$

- d.

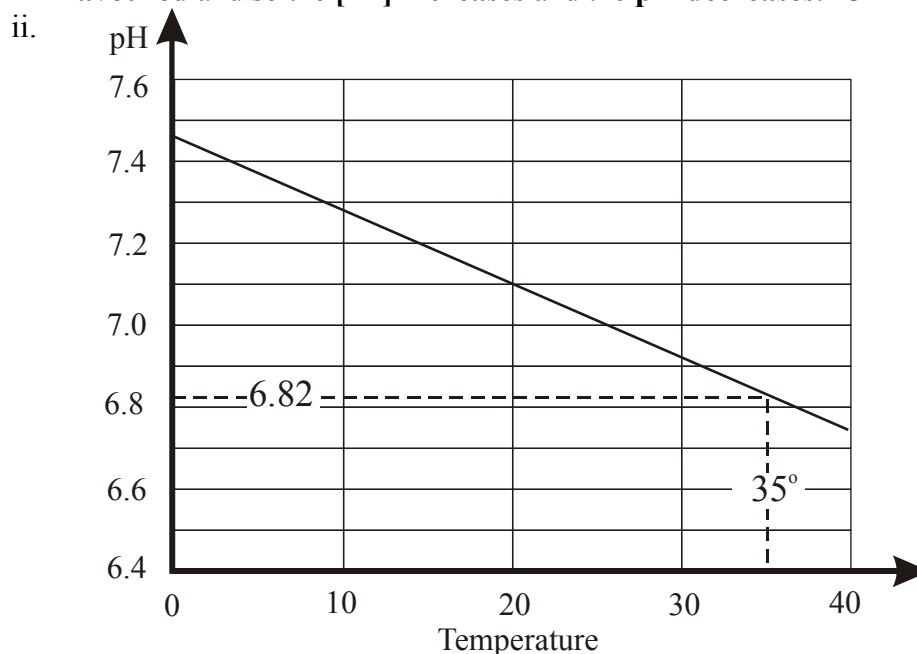


Question 4

- a. $Q = It = 2.5 \times 50 \times 60$
 $= 7.5 \times 10^3 \text{ C}$ ①
- b. $n(e^-) = Q / F = 7.5 \times 10^3 / 96500$
 $= 7.8 \times 10^{-2} \text{ mol}$ ①
- c. Whether or not a metal is deposited on the cathode during the electrolysis of an aqueous solution of a metal salt, depends on the relative strengths, as oxidants, of the metal cations and water.
According to the electrochemical series
 $2\text{H}_2\text{O(l)} + 2e^- \rightarrow \text{H}_2\text{(g)} + 2\text{OH}^-\text{(aq)} \quad -0.83 \text{ V}$
 $\text{K}^+\text{(aq)} + e^- \rightarrow \text{K(s)} \quad -2.93 \text{ V}$
 H_2O is a stronger oxidant than $\text{K}^+\text{(aq)}$ and will be preferentially reduced at the cathode, so no K will be deposited. ①
Since metal M is produced during the electrolysis of $\text{M(NO}_3)_2\text{(aq)}$, **$\text{M}^{2+}\text{(aq)}$ is a stronger oxidant than H_2O , and its reduction half-equation will be higher than that of H_2O on the electrochemical series.** ① i.e.
 $\text{M}^{2+}\text{(aq)} + 2e^- \rightarrow \text{M(s)} \quad > -0.83 \text{ V}$
 $2\text{H}_2\text{O(l)} + 2e^- \rightarrow \text{H}_2\text{(g)} + 2\text{OH}^-\text{(aq)} \quad -0.83 \text{ V}$
- d. Half-equations occurring in the KI(aq) cell are
Anode (+) $2\text{I}^-\text{(aq)} \rightarrow \text{I}_2\text{(s)} + 2e^-$
Cathode (-) $2\text{H}_2\text{O(l)} + 2e^- \rightarrow \text{H}_2\text{(g)} + 2\text{OH}^-\text{(aq)}$
 $n(\text{I}_2) \text{ formed} = \frac{1}{2} \times n(e^-)$
 $= \frac{1}{2} \times 7.8 \times 10^{-2}$
 $= 3.9 \times 10^{-2} \text{ mol}$ ①
 $m(\text{I}_2) \text{ formed} = 3.9 \times 10^{-2} \text{ mol} \times (2 \times 126.9)$
 $= 9.9 \text{ g}$ ①
- e. Reduction half-equation at cathode is $\text{M}^{2+}\text{(aq)} + 2e^- \rightarrow \text{M(s)}$
 $n(\text{M}) = \frac{1}{2} \times n(e^-)$
 $= \frac{1}{2} \times 7.8 \times 10^{-2}$
 $= 3.9 \times 10^{-2} \text{ mol}$ ①
 $\text{Molar mass of M} = m / n$
 $= 2.48 \text{ g} / 3.9 \times 10^{-2} \text{ mol}$
 $= 63.6 \text{ g mol}^{-1}$
Hence metal is **copper, Cu** ①

Question 5

- a. i. The graph shows that the pH of pure water decreases as the temperature increases.
This is consistent with the **exothermic** nature of the self-ionisation of water
 $2\text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$ ①
As the temperature of pure water increases, the **endothermic reverse reaction is favoured** and so the **$[\text{H}^+]$ increases** and the **pH decreases**. ①



At 35°C, pH = 6.82
 $[\text{H}_3\text{O}^+] = 10^{-6.82}$
 $= 1.5 \times 10^{-7} \text{ M}$ ①

Since pure water is neutral, whatever the temperature $[\text{OH}^-] = [\text{H}_3\text{O}^+]$, so
 $[\text{OH}^-] = 1.5 \times 10^{-7} \text{ M}$ ①

- b. Propanoic acid, $\text{C}_2\text{H}_5\text{COOH}$, is a weak acid $K_a = 1.3 \times 10^{-5}$ (Data Book – Table 13)
 $\text{C}_2\text{H}_5\text{COOH(aq)} \rightleftharpoons \text{C}_2\text{H}_5\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$
 $K_a = [\text{C}_2\text{H}_5\text{COO}^-][\text{H}^+] / [\text{C}_2\text{H}_5\text{COOH}]$
 Normal weak acids assumptions apply
 $[\text{C}_2\text{H}_5\text{COO}^-] = [\text{H}_3\text{O}^+]$ and $[\text{C}_2\text{H}_5\text{COOH}]_{\text{equilibrium}} = [\text{C}_2\text{H}_5\text{COOH}]_{\text{initially}} = 0.10 \text{ M}$
 So $1.3 \times 10^{-5} = [\text{H}_3\text{O}^+]^2 / 0.10$ ①
 $0.10 \times 1.3 \times 10^{-5} = [\text{H}_3\text{O}^+]^2$
 $[\text{H}_3\text{O}^+] = \sqrt{1.3 \times 10^{-6}}$
 $= 1.1 \times 10^{-3} \text{ M}$ ①
 $\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$
 $= -\log_{10} [1.1 \times 10^{-3}]$
 $= 2.9$ ①

- c. i. **HCl is a strong acid which ionises completely** according to
 $\text{HCl(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
0.10 M HCl(aq) $\rightarrow [\text{H}_3\text{O}^+] = 0.10 \text{ M} = 10^{-1} \text{ M}$ hence pH = 1 ①
 Diluting 10 mL of 0.10 M HCl(aq) to 100 mL with water gives
0.010 M HCl(aq) $\rightarrow [\text{H}_3\text{O}^+] = 0.010 \text{ M} = 10^{-2} \text{ M}$ hence pH = 2 ①
- ii. **Propanoic acid is a weak acid.**
 In part (b) it was shown that in **0.10 M C₂H₅COOH(aq), $[\text{H}_3\text{O}^+] = 1.1 \times 10^{-3} \text{ M}$**
 Diluting 10 mL of 0.10 M C₂H₅COOH(aq) to 100 mL with water gives
0.010 M C₂H₅COOH
 $K_a = [\text{C}_2\text{H}_5\text{COO}^-][\text{H}^+] / [\text{C}_2\text{H}_5\text{COOH}]$
 Normal weak acids assumptions apply
 $[\text{C}_2\text{H}_5\text{COO}^-] = [\text{H}_3\text{O}^+]$ and
 $[\text{C}_2\text{H}_5\text{COOH}]_{\text{equilibrium}} = [\text{C}_2\text{H}_5\text{COOH}]_{\text{initially}} = 0.010 \text{ M}$
 So $1.3 \times 10^{-5} = [\text{H}_3\text{O}^+]^2 / 0.010$ ①
 $0.010 \times 1.3 \times 10^{-5} = [\text{H}_3\text{O}^+]^2$
 $[\text{H}_3\text{O}^+] = \sqrt{(1.3 \times 10^{-7})}$
 $= 3.6 \times 10^{-4} \text{ M}$ ①
pH = $-\log_{10}[\text{H}_3\text{O}^+]$
 $= -\log_{10}(3.6 \times 10^{-4})$
 $= 3.4$ ①
- iii. As indicated in (i) **HCl(aq) is a strong acid** and so the **change in $[\text{H}_3\text{O}^+]$ is due only to the dilution factor**. Hence **increasing the volume by a factor of 10, decreases the $[\text{H}_3\text{O}^+]$ by a factor of 10 and thus increases the pH by 1**. ①
Propanoic acid, C₂H₅COOH, is a weak acid, ionising according to
 $\text{C}_2\text{H}_5\text{COOH(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{C}_2\text{H}_5\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$
 When the volume is increased by a factor of 10, the $[\text{H}_3\text{O}^+]$ immediately decreases by a factor of 10. **However, addition of more of the reactant H₂O(l) pushes the position of equilibrium to the right, so increasing the $[\text{H}_3\text{O}^+]$** .
The combined effect of the dilution and the movement of the position of equilibrium is that $[\text{H}_3\text{O}^+]$ decreases but by a factor less than 10. ①
 Consequently the **change in pH** – it increases by less than 1 – is **smaller than that for HCl(aq)**. ①

Question 6

- a. **1 M H⁺(aq) – e.g. 1 M HCl(aq), Pt electrode** ①
H₂(g) at 101.3 kPa, 25°C ①
 A standard hydrogen half-cell, also known as the standard hydrogen electrode, contains a platinum electrode with hydrogen gas at 101.3 kPa, bubbled through a solution which is 1 M with respect to H⁺(aq) at 25°C.
- b. According to the electrochemical series
 $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq}) \quad 1.36 \text{ V}$
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \quad 0.0 \text{ V}$
 The strongest oxidant Cl₂ reacts with strongest reductant H⁺ as described by the overall equation
H₂(g) + Cl₂(g) $\rightarrow 2\text{H}^+(\text{aq}) + 2\text{Cl}^-(\text{aq})$ ①

- c. The decrease in pH indicates that the $[\text{H}^+]$ **increases** in the standard hydrogen half-cell.
 Half-reaction occurring must be $\text{H}_2(\text{g}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{e}^-$ ❶
 Since **oxidation** is occurring in the half-cell, and it is part of a galvanic cell, the **electrode must be negative (-)** ❶
- d. In the electrochemical series based on the standard hydrogen half-cell, the relevant half-equations are
 $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s}) \quad 0.34 \text{ V}$
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \quad 0.00 \text{ V}$
 $\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s}) \quad -0.23 \text{ V}$
 In the alternative electrochemical series, the **differences between the standard electrode potentials will be same** – the relative strengths of oxidants and reductants remains the same – but the **standard electrode potential values change** because the **Cu^{2+}/Cu half-cell is set at 0.0.** ❶
 On the alternative electrochemical series, the half-equations would appear as
 $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s}) \quad 0.00 \text{ V}$
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \quad -0.34 \text{ V}$
 $\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s}) \quad \underline{-0.57 \text{ V}}$ ❶
- e. According to the electrochemical series
 $2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq}) \quad -0.83 \text{ V}$
 $\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Al}(\text{s}) \quad -1.67 \text{ V}$
 H_2O is a stronger oxidant ❶ than $\text{Al}^{3+}(\text{aq})$ and will be preferentially reduced ❶ during electrolysis

End of Suggested Answers