

Please write clearly in block capitals.

Centre number

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Candidate number

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Surname *StemUp Model Solution*

Forename(s)

Candidate signature

I declare this is my own work.

INTERNATIONAL A-LEVEL CHEMISTRY (9620)

Unit 3: Inorganic 2 and Physical 2

Tuesday 19 January 2021 07:00 GMT Time allowed: 1 hour 30 minutes

Materials

For this paper you must have:

- the Periodic Table/Data Sheet, provided as an insert
- a ruler with millimetre measurements
- a scientific calculator, which you are expected to use where appropriate.

Instructions

- Use black ink or black ball-point pen.
- Fill in the boxes at the top of this page.
- Answer **all** questions.
- You must answer the questions in the spaces provided. Do not write outside the box around each page or on blank pages.
- If you need extra space for your answer(s), use the lined pages at the end of this book. Write the question number against your answer(s).
- All working must be shown.
- Do all rough work in this book. Cross through any work you do not want to be marked.

Information

- The marks for questions are shown in brackets.
- The maximum mark for this paper is 80.

For Examiner's Use	
Question	Mark
1	
2	
3	
4	
5	
6	
7	
8	
9	
10	
TOTAL	



J A N 2 1 C H 0 3 0 1

Answer **all** questions in the spaces provided.

0 1

This question is about Period 3 elements and their compounds.

0 1 . 1

State **two** observations that you would make when magnesium is burned in oxygen.

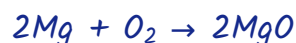
Write an equation for the reaction.

[3 marks]

Observation 1 (bright) white light / white flame

Observation 2 white ash / white powder (white solid) formed

Equation



0 1 . 2

A strip of magnesium is added to some cold water. At first, there is no visible change but magnesium does react very slowly with cold water.

Write an equation for the reaction.

Explain how an indicator can be used to show that a reaction is occurring.

[3 marks]

Equation



Explanation

add phenolphthalein: it turns pink because $\text{Mg}(\text{OH})_2$ is a weak alkali

(slightly soluble, so OH^- ions are present, pH 9 to 10)

0 1 . 3

Write an equation for the reaction of aluminium oxide with aqueous potassium hydroxide.

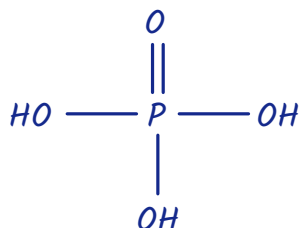
[1 mark]



0 1 . 4

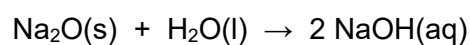
Draw the structure of the acid formed when phosphorus(V) oxide reacts with water.

[1 mark]



0 1 . 5

A 0.589 g sample of solid sodium oxide is reacted with water to form 50.0 cm³ of sodium hydroxide solution.



Calculate the pH, at 25 °C, of this solution.
Give your answer to 2 decimal places.

At 25 °C, the ionic product of water, $K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$

[5 marks]

$$n(\text{Na}_2\text{O}) = 0.589 \div 62.0 = 9.50 \times 10^{-3} \text{ mol}$$

$$n(\text{NaOH}) = 2 \times 9.50 \times 10^{-3} = 0.0190 \text{ mol}$$

$$[\text{NaOH}] = 0.0190 \times 1000 \div 50.0 = 0.380 \text{ mol dm}^{-3}$$

$$[\text{OH}^-] = 0.380 \text{ mol dm}^{-3}$$

$$[\text{H}^+] = K_w \div [\text{OH}^-] = 1.00 \times 10^{-14} \div 0.380$$

$$[\text{H}^+] = 2.63 \times 10^{-14} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(2.63 \times 10^{-14}) = 13.58$$

pH 13.58

13

Turn over ►



0 2

This question is about copper complexes.

0 2 . 1

Copper forms a complex ion $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$

State what is meant by the term complex ion.

[1 mark]

*a central metal ion surrounded by ligands that are bonded to it by
co-ordinate bonds*

0 2 . 2

A solution containing a copper(II) complex absorbs light with a wavelength of 593 nm

Calculate the energy difference (ΔE), in J, between the ground state and the excited state of d-electrons in this complex.

You should assume that all the light absorbed is because of the movement of electrons between d orbitals.

Planck's constant, $h = 6.626 \times 10^{-34} \text{ J s}$

Speed of light, $c = 2.998 \times 10^8 \text{ m s}^{-1}$

[3 marks]

$$\Delta E = hc \div \lambda$$

$$\lambda = 593 \text{ nm} = 593 \times 10^{-9} \text{ m}$$

$$\Delta E = (6.626 \times 10^{-34} \times 2.998 \times 10^8) \div (593 \times 10^{-9})$$

$$\Delta E = 3.35 \times 10^{-19} \text{ J}$$

$$\Delta E \quad 3.35 \times 10^{-19} \quad \text{J}$$



0 2 . 3

When an excess of aqueous ammonia is added to a solution containing aqueous copper(II) ions, a different copper complex is formed.

Name this type of reaction.

Write an equation for this reaction.

[2 marks]

Type of reaction (ligand) substitution / ligand exchange

Equation

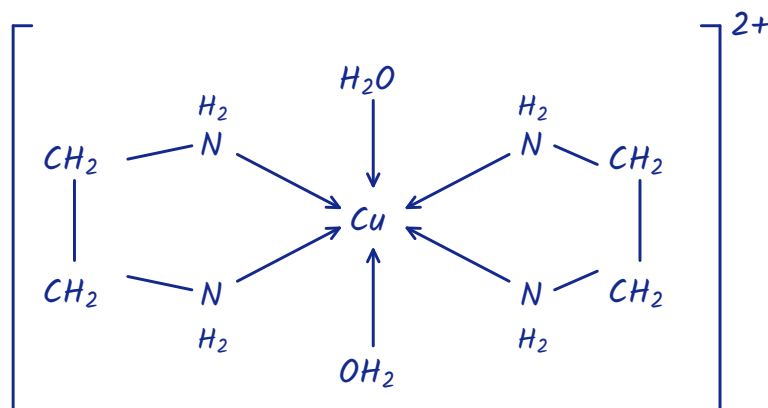


0 2 . 4

Copper forms a complex with the formula $[\text{Cu}(\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2)_2(\text{H}_2\text{O})_2]^{2+}$

Draw the structure of the *trans* isomer of this complex.

[2 marks]



trans: the two H₂O ligands are opposite each other

Turn over for the next question

8

Turn over ►



0 3

This question is about some reactions of iron salts.

A student investigates the reactions of an aqueous solution of an iron(II) salt.

Table 1 shows the observations from some test-tube reactions.

Table 1

		Observations
Test 1	Add acidified $\text{BaCl}_2(\text{aq})$	White precipitate formed
Test 2	Add $\text{NaOH}(\text{aq})$ until in excess	Green precipitate formed that turns brown on standing in air
Test 3	Add $\text{Na}_2\text{CO}_3(\text{aq})$	Green precipitate formed

0 3 . 1

Use the observation in **Test 1** to identify the anion in the iron(II) salt.

[1 mark]

sulfate ion, SO_4^{2-}

0 3 . 2

Give the formula of the green precipitate formed in **Test 2**.

State the type of reaction that occurs when the green precipitate turns brown in **Test 2**.

[2 marks]

Formula $\text{Fe}(\text{H}_2\text{O})_4(\text{OH})_2$

Type of reaction oxidation (redox)

0 3 . 3

Write an ionic equation for the reaction occurring in **Test 3**.

[1 mark]

$[\text{Fe}(\text{H}_2\text{O})_6]^{2+} + \text{CO}_3^{2-} \rightarrow \text{FeCO}_3 + 6\text{H}_2\text{O}$



0 3 . 4

The student repeats **Test 2** and **Test 3** using aqueous iron(III) chloride.

Complete **Table 2** to describe what is observed in **Test 2** and **Test 3** with iron(III) chloride.

[3 marks]

Table 2

		Observations with iron(III) chloride
Test 2	Add NaOH(aq) until in excess	<i>(orange-)brown precipitate</i>
Test 3	Add Na ₂ CO ₃ (aq)	<i>(orange-)brown precipitate and effervescence (fizzing, bubbles of CO₂ gas)</i>

0 3 . 5

In a separate test-tube reaction, the student adds an excess of concentrated hydrochloric acid to the iron(III) chloride solution.

What is the shape of the complex ion formed in this reaction?

Tick (✓) **one** box

[1 mark]

Octahedral

☐

Square Planar

☐

Tetrahedral

☒

Question 3 continues on the next page

Turn over ►



0 3 . 6

The reaction between I^- ions and $S_2O_8^{2-}$ ions is slow but is much faster when iron(II) ions are added as a catalyst.

Explain why the reaction between I^- ions and $S_2O_8^{2-}$ ions is slow.

Explain how iron(II) ions catalyse the reaction.
Use suitable equations in your answer.

[5 marks]

Why reaction is slow the I^- ions and $S_2O_8^{2-}$ ions are both negative, so
they repel each other
therefore the reaction has a high activation energy, so few collisions
lead to reaction

How iron(II) ions catalyse reaction



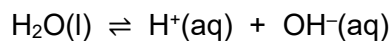
in each step the ions are oppositely charged (positive Fe ion with a
negative ion), so they attract, which lowers the activation energy; Fe^{2+}
is regenerated

13



0 4

Water dissociates as shown

The ionic product of water (K_w) is given by the expression

$$K_w = [\text{H}^+][\text{OH}^-]$$

0 4 . 1

State why the concentration of water is **not** shown in the expression for K_w

[1 mark]

$[\text{H}_2\text{O}]$ is very large so it is (approximately) constant

0 4 . 2

The dissociation of water is an endothermic process.

Explain why the value for K_w increases as the temperature increases.

[2 marks]

the equilibrium shifts to the right (endothermic direction) to oppose the increase in temperature

so $[\text{H}^+]$ and $[\text{OH}^-]$ both increase, so K_w increases

0 4 . 3

At 45 °C, $K_w = 4.02 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$

Calculate the pH of water at 45 °C

Give your answer to 2 decimal places.

[2 marks]

$$[\text{H}^+] = \sqrt{K_w} = \sqrt{(4.02 \times 10^{-14})}$$

$$[\text{H}^+] = 2.00 \times 10^{-7} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(2.00 \times 10^{-7}) = 6.70$$

pH 6.70

5

Turn over ►



0 5

This question is about weak acids.

0 5

1

A $0.150 \text{ mol dm}^{-3}$ aqueous solution of propanoic acid ($\text{CH}_3\text{CH}_2\text{COOH}$) has a pH of 2.85 at 25°C

Calculate a value, in mol dm^{-3} , for the acid dissociation constant (K_a) of propanoic acid at 25°C

[3 marks]

$$[\text{H}^+] = 10^{-2.85} = 1.41 \times 10^{-3} \text{ mol dm}^{-3}$$

$$K_a = [\text{H}^+][\text{CH}_3\text{CH}_2\text{COO}^-] \div [\text{CH}_3\text{CH}_2\text{COOH}] = [\text{H}^+]^2 \div [\text{HA}]$$

$$K_a = (1.41 \times 10^{-3})^2 \div 0.150 = 1.33 \times 10^{-5}$$

$$K_a \underline{1.33 \times 10^{-5}} \text{ mol dm}^{-3}$$

0 5

2

A buffer solution is prepared by mixing 25.0 cm^3 of $0.120 \text{ mol dm}^{-3}$ aqueous sodium butanoate with 35.0 cm^3 of $0.150 \text{ mol dm}^{-3}$ aqueous butanoic acid at 25°C

Calculate the pH of the buffer solution formed.
Give your answer to 2 decimal places.

For butanoic acid, the acid dissociation constant, $K_a = 1.51 \times 10^{-5} \text{ mol dm}^{-3}$ at 25°C

[5 marks]

$$n(\text{sodium butanoate}) = 25.0 \times 10^{-3} \times 0.120 = 3.00 \times 10^{-3} \text{ mol}$$

$$n(\text{butanoic acid}) = 35.0 \times 10^{-3} \times 0.150 = 5.25 \times 10^{-3} \text{ mol}$$

$$[\text{H}^+] = K_a[\text{HA}] \div [\text{A}^-]$$

$$[\text{H}^+] = 1.51 \times 10^{-5} \times 5.25 \times 10^{-3} \div 3.00 \times 10^{-3}$$

$$[\text{H}^+] = 2.64 \times 10^{-5} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(2.64 \times 10^{-5}) = 4.58$$

$$\text{pH} \underline{4.58}$$



0 5 . 3

Explain why the pH of the buffer solution formed in Question 05.2 remains approximately constant when a small amount of sodium hydroxide solution is added.

[2 marks]

the added OH^- ions react with the butanoic acid: $\text{HA} + \text{OH}^- \rightarrow \text{A}^- +$

H_2O ($\text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{HA} + \text{OH}^-$ shifts left to remove the OH^-)

so the ratio $[\text{HA}]/[\text{A}^-]$ and therefore $[\text{H}^+]$ stay approximately constant

10

Turn over for the next question

Turn over ►



0 6

This question is about two chlorides of potassium, KCl_2 and KCl

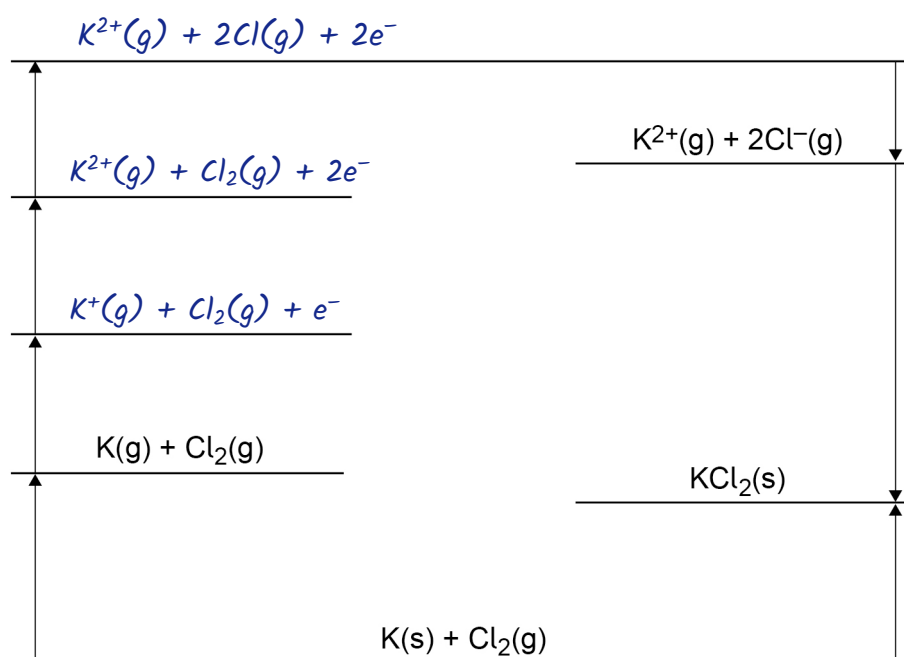
KCl_2 is a theoretical compound that has never been isolated.

KCl_2 contains the K^{2+} ion.

Born-Haber cycles can be used to determine the enthalpy of formation of a theoretical compound.

Figure 1 shows an incomplete Born-Haber cycle for the formation of KCl_2 including the formation of the K^{2+} ion. (The diagram is not drawn to scale.)

Figure 1



0 6

1

Complete the Born-Haber cycle for the formation of KCl_2

Write the correct species, including state symbols, on the three blank lines.

[3 marks]



0 6 . 2

Use the values in **Table 3** to calculate a value for the enthalpy of formation of KCl_2 **[2 marks]****Table 3**

	$\Delta H^\ominus / \text{kJ mol}^{-1}$
$\text{K(s)} \rightarrow \text{K(g)}$	+90
$\text{K(g)} \rightarrow \text{K}^+(\text{g}) + \text{e}^-$	+418
$\text{K}^+(\text{g}) \rightarrow \text{K}^{2+}(\text{g}) + \text{e}^-$	+3070
$\text{Cl}_2(\text{g}) \rightarrow 2\text{Cl(g)}$	+242
$\text{Cl(g)} + \text{e}^- \rightarrow \text{Cl}^-(\text{g})$	-364
$\text{K}^{2+}(\text{g}) + 2\text{Cl}^-(\text{g}) \rightarrow \text{KCl}_2(\text{s})$	-2350

$$\Delta_f H = \Delta_{at} H(\text{K}) + 1E_1 + 1E_2 + E(\text{Cl-Cl}) + 2EA + LE$$

$$\Delta_f H = 90 + 418 + 3070 + 242 + (2 \times -364) + (-2350)$$

$$\Delta_f H = +742 \text{ kJ mol}^{-1}$$

Enthalpy of formation +742 kJ mol⁻¹**Question 6 continues on the next page****Turn over ►**

0 6 . 3

The other chloride, KCl, is very soluble in water.

Use the values in **Table 4** to calculate a value for the enthalpy of solution of KCl

[2 marks]

Table 4

	$\Delta H^\ominus / \text{kJ mol}^{-1}$
$\text{K}^+(\text{g}) + \text{Cl}^-(\text{g}) \rightarrow \text{KCl}(\text{s})$	-701
$\text{K}^+(\text{g}) \rightarrow \text{K}^+(\text{aq})$	-322
$\text{Cl}^-(\text{g}) \rightarrow \text{Cl}^-(\text{aq})$	-364

$$\Delta H_{\text{sol}} = \Delta H_{\text{lattice dissociation}}(\text{KCl}) + \Delta H_{\text{hyd}}(\text{K}^+) + \Delta H_{\text{hyd}}(\text{Cl}^-)$$

$$\Delta H_{\text{sol}} = +701 + (-322) + (-364)$$

$$\Delta H_{\text{sol}} = +15 \text{ kJ mol}^{-1}$$

Enthalpy of solution +15 kJ mol⁻¹

0 6 . 4

How does the value of the hydration enthalpy of K²⁺ compare with the value of the hydration enthalpy of K⁺?

Tick (✓) **one** box.

[1 mark]

More exothermic than K⁺

☒

The same as K⁺

☐

Less exothermic than K⁺

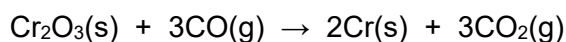
☐

8



07

Chromium(III) oxide can be reduced to chromium using carbon monoxide.



07.1

Table 5 shows some thermodynamic data.

Use the values in **Table 5** to calculate the free energy change ($\Delta G^\ominus$) for this reaction at 25 °C
Show your working.

[6 marks]**Table 5**

	$\text{Cr}_2\text{O}_3(\text{s})$	$\text{CO}(\text{g})$	$\text{Cr}(\text{s})$	$\text{CO}_2(\text{g})$
$\Delta_f H^\ominus / \text{kJ mol}^{-1}$	-1128	-111	0	-394
$S^\ominus / \text{J K}^{-1} \text{mol}^{-1}$	81.2	198	23.8	214

$$\Delta H = [0 + (3 \times -394)] - [-1128 + (3 \times -111)]$$

$$\Delta H = -1182 - (-1461) = +279 \text{ kJ mol}^{-1}$$

$$\Delta S = [(2 \times 23.8) + (3 \times 214)] - [81.2 + (3 \times 198)]$$

$$\Delta S = 689.6 - 675.2 = +14.4 \text{ J K}^{-1} \text{mol}^{-1}$$

$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G = 279 - [298 \times (14.4 \div 1000)]$$

$$\Delta G = 279 - 4.29 = +274.7 \text{ kJ mol}^{-1}$$

$$\Delta G^\ominus \quad \underline{+274.7} \quad \text{kJ mol}^{-1}$$

07.2

Explain the significance of your answer to Question **07.1**

(If you were unable to answer Question **07.1**, you should use the value for $\Delta G^\ominus$ of +150 kJ mol⁻¹. This is **not** the correct answer.)

[1 mark]

ΔG is positive, so the reaction is not feasible at 25 °C

7

Turn over ►

0 8

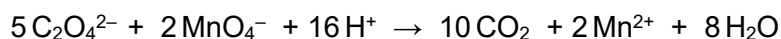
This question is about a metal ethanedioate salt, $M_2C_2O_4$, where **M** is an unknown metal.

$M_2C_2O_4$ is analysed by titration with a standard solution of potassium manganate(VII) in acidic conditions. The metal ion does not react with manganate(VII) ions.

2.07 g of $M_2C_2O_4$ are dissolved in distilled water. The solution is made up to 250 cm^3

25.0 cm^3 samples of this solution are titrated with acidified potassium manganate(VII) solution.

26.50 cm^3 of $0.0188\text{ mol dm}^{-3}$ acidified potassium manganate(VII) solution are required for a complete reaction.



Calculate the relative atomic mass of the unknown metal, **M**.

Identify **M**.

[6 marks]

$$n(\text{MnO}_4^-) = 26.50 \times 0.0188 \div 1000 = 4.982 \times 10^{-4} \text{ mol}$$

$$n(\text{C}_2\text{O}_4^{2-}) \text{ in } 25.0\text{ cm}^3 = 4.982 \times 10^{-4} \times 5/2 = 1.2455 \times 10^{-3} \text{ mol}$$

$$n(\text{C}_2\text{O}_4^{2-}) \text{ in } 250\text{ cm}^3 = 1.2455 \times 10^{-3} \times 10 = 1.2455 \times 10^{-2} \text{ mol}$$

$$M_r \text{ of } M_2C_2O_4 = 2.07 \div 1.2455 \times 10^{-2} = 166.2$$

$$M_r \text{ of } C_2O_4 = (2 \times 12.0) + (4 \times 16.0) = 88.0$$

$$2 \times A_r(M) = 166.2 - 88.0 = 78.2$$

$$A_r(M) = 39.1 \text{ so } M \text{ is potassium}$$

Relative atomic mass 39.1

Metal **M** potassium (K)

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6



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09

Table 6 shows some electrode potential data.

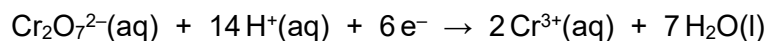
Table 6

Electrode half-equation	$E^{\ominus}/V$
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Fe}(\text{s})$	-0.44
$\text{Fe}^{3+}(\text{aq}) + \text{e}^{-} \rightarrow \text{Fe}^{2+}(\text{aq})$	0.77
$\text{Br}_2(\text{g}) + 2\text{e}^{-} \rightarrow 2\text{Br}^{-}(\text{aq})$	1.07
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^{+}(\text{aq}) + 6\text{e}^{-} \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	1.33
$\text{Cl}_2(\text{g}) + 2\text{e}^{-} \rightarrow 2\text{Cl}^{-}(\text{aq})$	1.36
$\text{F}_2(\text{g}) + 2\text{e}^{-} \rightarrow 2\text{F}^{-}(\text{aq})$	2.87

09

1

A platinum electrode is needed when measuring the standard electrode potential for the half-equation



Give **two** reasons why platinum is used.

[2 marks]

1 *platinum is inert (unreactive)*

2 *it conducts electricity, allowing transfer of electrons*

09

2

Identify the strongest reducing agent in **Table 6**.

[1 mark]

Fe (iron)



0 9 . 3

An electrochemical cell uses two platinum electrodes.

One platinum electrode is placed in an aqueous solution containing Fe^{3+} and Fe^{2+} ions and the other in an aqueous solution containing $\text{Cr}_2\text{O}_7^{2-}$, H^+ and Cr^{3+} ions.

Give the conventional representation of this cell.

Use data from **Table 6** to calculate the cell EMF.

[3 marks]

Conventional representation



Calculation

$$\text{EMF} = E_{\text{right}} - E_{\text{left}} = 1.33 - 0.77 = +0.56 \text{ V}$$

EMF +0.56 V

6

Turn over for the next question

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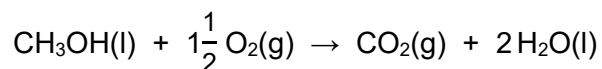


1 0

Methanol has been proposed as the fuel in a fuel cell instead of hydrogen.

1 0 . 1

The overall cell reaction in the methanol fuel cell is shown.



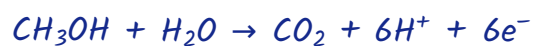
In this fuel cell, in acidic conditions:

- the methanol is oxidised to form carbon dioxide
- the oxygen is reduced to form water.

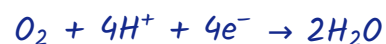
Deduce the oxidation half-equation and the reduction half-equation for this cell when operating in acidic conditions.

[2 marks]

Oxidation half-equation



Reduction half-equation



1 0 . 2

Suggest **one** advantage of using methanol and **one** advantage of using hydrogen in a fuel cell.

[2 marks]

Advantage of methanol methanol is a liquid, so it is more easily (and safely) stored and transported

Advantage of hydrogen no carbon dioxide is produced in the cell, only water

4

END OF QUESTIONS



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