OXFORD AQA
INTERNATIONAL QUALIFICATIONS

Please write clearly in block capitals.

Centre number | | | | | | Candidate number | | | | |

Surname ______________________________

Forename(s) ______________________________

Candidate signature ______________________________
I declare this is my own work.

INTERNATIONAL A-LEVEL
CHEMISTRY (9620)

Unit 3: Inorganic 2 and Physical 2

Wednesday 5 June 2024 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	
TOTAL	

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

0 1 This question is about the enthalpy of lattice dissociation of calcium bromide, $CaBr_2$

0 1 . 1 Define enthalpy of lattice dissociation.

[2 marks]

Table 1 shows some enthalpy change data.

Table 1

	ΔH / kJ mol^{-1}
Enthalpy of solution of $CaBr_2$	–110
Enthalpy of hydration of Ca^{2+}	–1650
Enthalpy of hydration of Br^-	–335

0 1 . 2 Use the data in **Table 1** to calculate the enthalpy of lattice dissociation, in kJ mol^{-1}, for calcium bromide.

[2 marks]

Enthalpy of lattice dissociation ____________________ kJ mol^{-1}

0 1 . 3 The enthalpy of lattice dissociation for calcium bromide calculated in Question **01.2** is the experimental value.

Explain why this experimental value is different from the theoretical value for the enthalpy of lattice dissociation of calcium bromide.

[1 mark]

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0 1 . 4 A student does an experiment to determine the enthalpy of solution for calcium bromide.

1.25 g of solid calcium bromide is added from a weighing boat to 25.0 cm^3 distilled water in a plastic (polystyrene) cup and the mixture is stirred gently.

When the calcium bromide dissolves, the temperature of the mixture rises from 19.6 °C to 24.9 °C

Calculate the enthalpy of solution, in kJ mol^{-1}, for calcium bromide.

Assume that the specific heat capacity of the solution is 4.18 J K^{-1} g^{-1}
Assume that the density of the solution is 1.00 g cm^{-3}

[3 marks]

Enthalpy of solution ______________________ kJ mol^{-1}

0 1 . 5 Another student repeats the experiment.

After the calcium bromide is added into the plastic cup, the student washes the weighing boat with a small volume of distilled water.

The student adds the washings to the plastic cup to ensure that all the calcium bromide is transferred.

Explain why adding the washings causes inaccurate results.

[1 mark]

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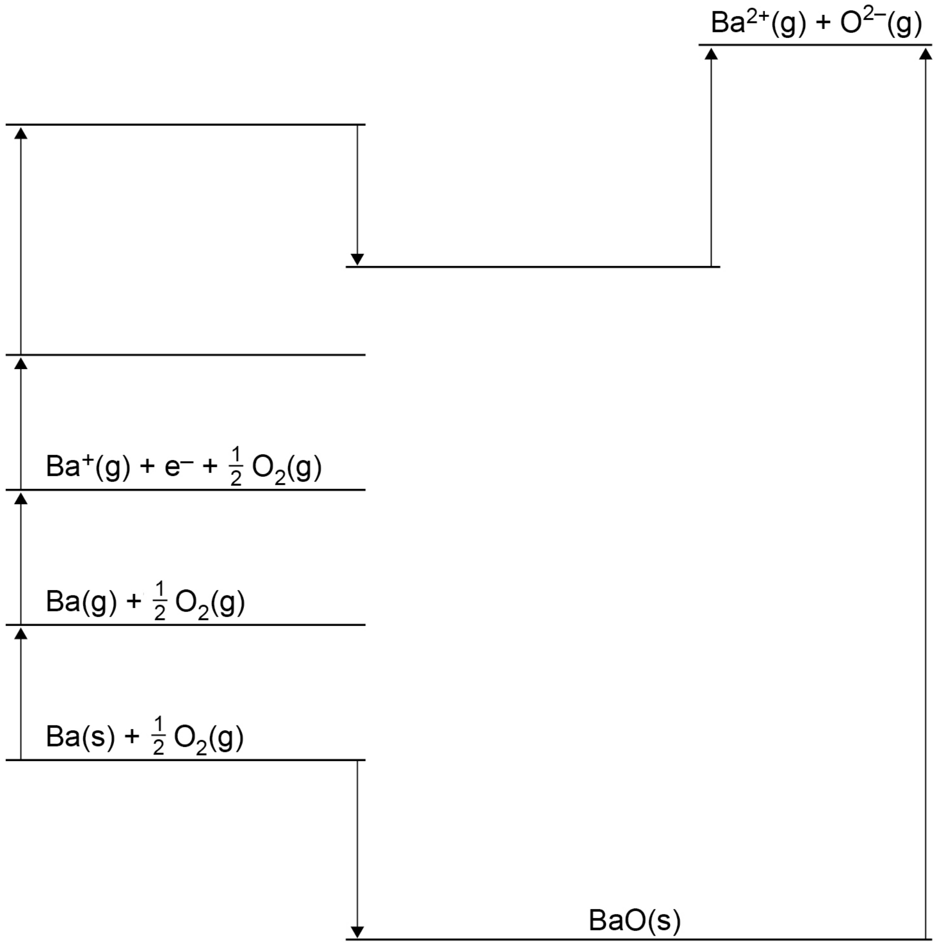
9

0 2 **Figure 1** shows an incomplete Born–Haber cycle for barium oxide, BaO

Figure 1



0 2 . 1 Write formulas, including state symbols, of the appropriate species on each of the three blank horizontal lines in **Figure 1**.

[3 marks]

Do not write outside the box

Table 2 shows some enthalpy change data.

Table 2

	Δ*H* / kJ mol^{-1}
Enthalpy of atomisation of barium	+176
Enthalpy of atomisation of oxygen	+248
First ionisation enthalpy of barium	+502
Second ionisation enthalpy of barium	+966
First electron affinity of oxygen	–142
Second electron affinity of oxygen	+844
Lattice dissociation enthalpy of barium oxide	+3152

0 2 . 2 Use **Figure 1** and **Table 2** to calculate the enthalpy of formation, in kJ mol^{-1}, of barium oxide.

[2 marks]

Enthalpy of formation ____________________ kJ mol^{-1}

5

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0 3 Magnesium reacts with carbon dioxide.

$$2\,Mg(s) + CO_2(g) \rightarrow 2\,MgO(s) + C(s)$$

Table 3 shows how the value for the Gibbs free-energy change (ΔG) for this reaction changes with temperature.

Table 3

Temperature / K	**ΔG / kJ mol^{-1}**
200	–165
400	–122
500	–101
700	–58
900	–15

0 3 . 1 Use the data in **Table 3** to plot a graph on **Figure 2** on **page 7** of ΔG (y-axis) against temperature.

Draw a line of best fit.

[2 marks]

0 3 . 2 Use your graph on **Figure 2** on **page 7** to deduce the enthalpy change, in kJ mol^{-1}, for the reaction.

[1 mark]

Enthalpy change ____________________ kJ mol^{-1}

0 3 . 3 Use your graph on **Figure 2** on **page 7** to determine the entropy change, in J K^{-1} mol^{-1}, for the reaction.

[2 marks]

Entropy change ____________________ J K^{-1} mol^{-1}

0 3 . 4 The gradient of the line on **Figure 2** changes at 923 K
Suggest why.

[1 mark]

____

____

6

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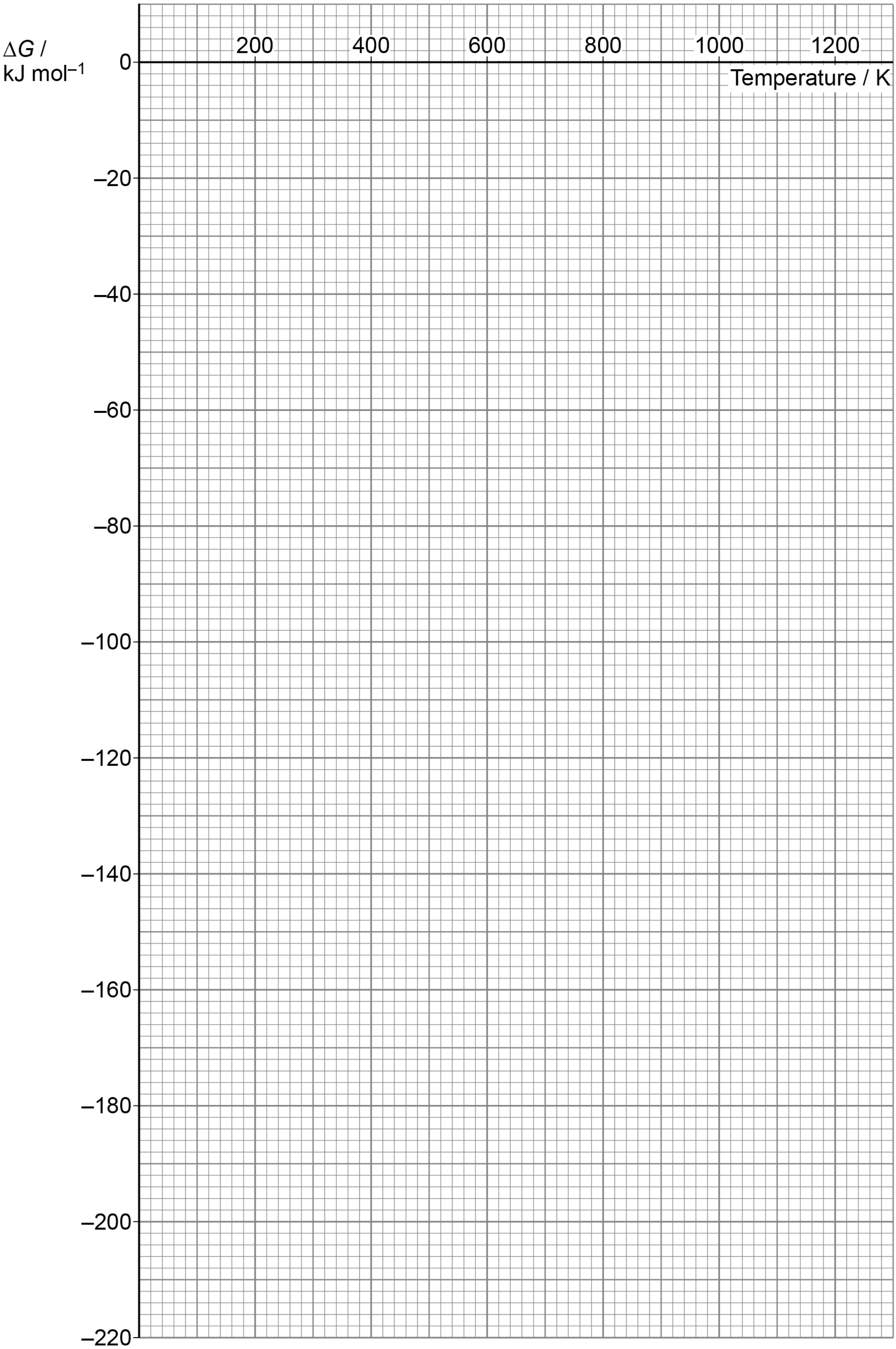
Figure 2



Use your completed **Figure 2** to answer Questions **03.2** and **03.3** on **page 6**.

Turn over ►

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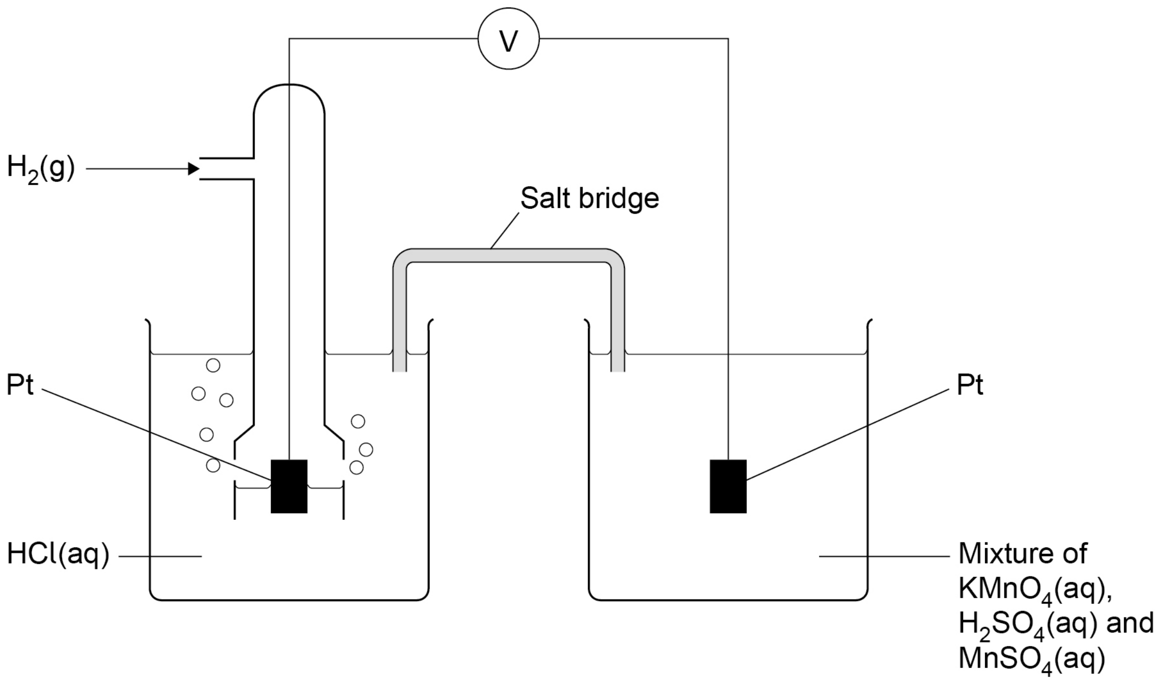
0 4 **Figure 3** shows the electrochemical cell used to measure the standard electrode potential for the MnO_4^- / Mn^{2+} electrode.

In this electrochemical cell, the standard hydrogen electrode is shown on the left-hand side.

Figure 3



0 4 . 1 State **three** conditions needed for the standard hydrogen electrode.

[2 marks]

0 4 . 2 Give the conventional cell representation for this electrochemical cell.

[2 marks]

Do not write outside the box

Table 4 shows some electrode potential data.

Table 4

	$E^\ominus$ / V
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.52
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Na^+(aq) + e^- \rightarrow Na(s)$	–2.71

0 4 . 3 Use data from **Table 4** to explain why sodium chloride solution should **not** be used in the salt bridge in **Figure 3**.

[2 marks]

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____

0 4 . 4 The standard hydrogen electrode is replaced by a Fe^{3+} / Fe^{2+} electrode.

The reading on the voltmeter is 0.75 V

Calculate the standard electrode potential of the Fe^{3+} / Fe^{2+} electrode.

[1 mark]

$E^\ominus$ ______________________ V

Question 4 continues on the next page

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Lithium is used to make rechargeable batteries.
The half-equations at each of the electrodes in one type of lithium cell are:

Positive Electrode	$Li^{+} + MnO_2 + e^{-} \rightarrow LiMnO_2$	$E^{\ominus} = -0.15$ V
Negative Electrode	$Li^{+} + e^{-} \rightarrow Li$	$E^{\ominus} = -3.05$ V

0 4 . 5 Deduce the overall equation for the reaction when this cell is being **recharged**.

[1 mark]

____

0 4 . 6 Suggest why there must be no water in this cell.

[1 mark]

____

____

9

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0 5 This question is about acids and bases.

0 5 . 1 Calculate the pH of a 0.150 mol dm^{-3} solution of hydrochloric acid.
Give your answer to 2 decimal places.

[1 mark]

pH ____________________

Water dissociates as shown

$$H_2O \rightleftharpoons H^+ + OH^-$$

0 5 . 2 Give an expression for K_w

[1 mark]

K_w

0 5 . 3 Calculate the pH of a 0.250 mol dm^{-3} aqueous solution of sodium hydroxide at 30 °C

The value of K_w is 1.47×10^{-14} mol^2 dm^{-6} at 30 °C
Give your answer to 2 decimal places.

[2 marks]

pH ____________________

0 5 . 4 When the aqueous solution of sodium hydroxide in Question **05.3** is heated, the pH of the solution decreases.

Explain why.

[2 marks]

____

____

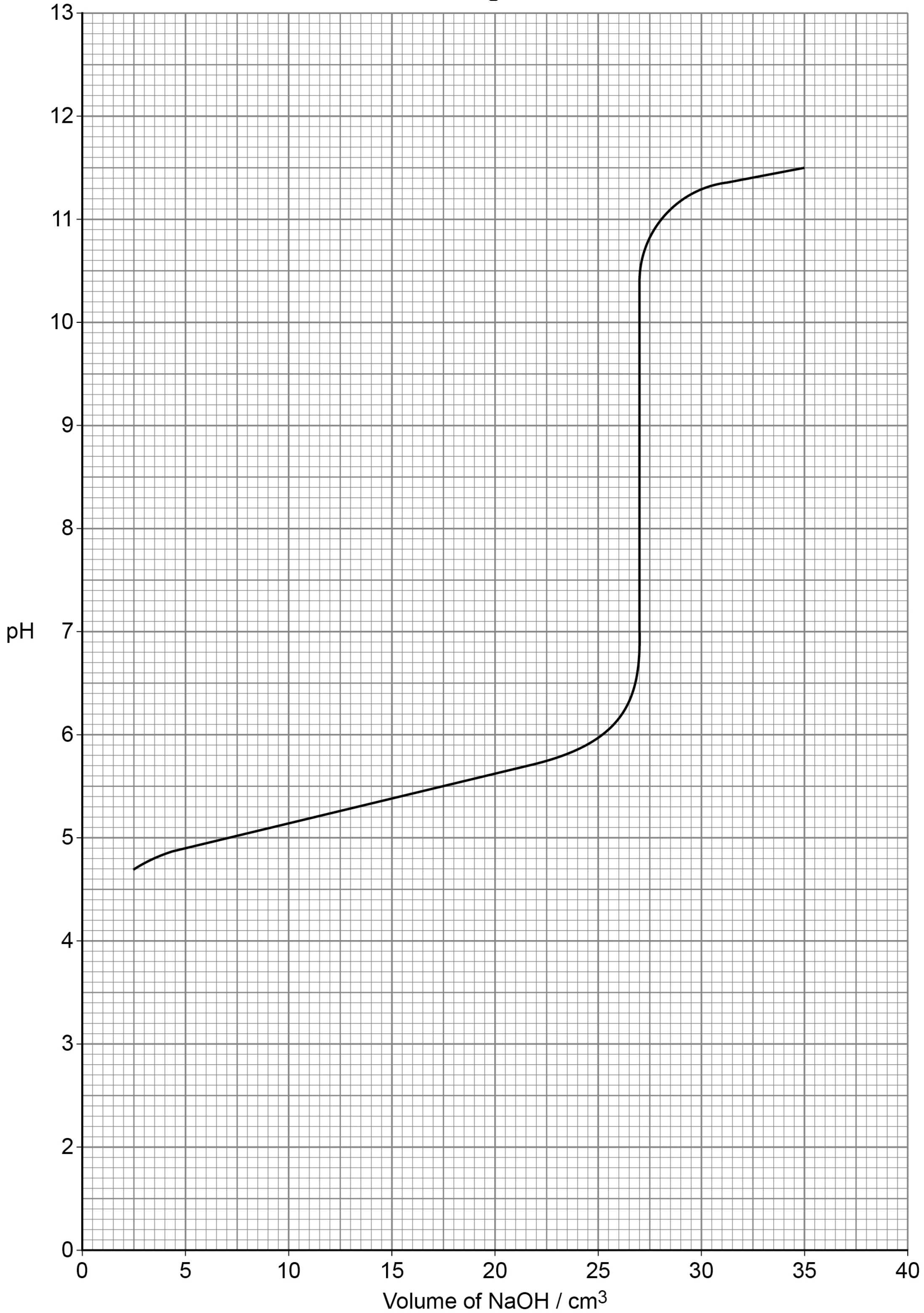
____

____

A student adds 0.100 mol dm^{-3} aqueous sodium hydroxide from a burette to 25.0 cm^3 of an aqueous solution of the weak acid HX

Figure 4 shows how the pH changes during part of this titration at 298 K

Figure 4



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0 5 . 5 Use **Figure 4** to determine the concentration, in mol dm^{-3}, of the weak acid, HX

[3 marks]

Concentration ______________________ mol dm^{-3}

0 5 . 6 Use **Figure 4** to determine the pH of the mixture at half equivalence (half neutralisation).

Use this pH to determine the value of the acid dissociation constant, K_a, at 298 K for the weak acid HX

Give the units for K_a

[3 marks]

pH at half equivalence ______________________________

Value of K_a ______________________________

Units of K_a ______________________________

12

Turn over for the next question

0 6 This question is about metal ions and complexes.

0 6 . 1 Give the full electron configuration of Fe^{2+}

[1 mark]

____

0 6 . 2 In the haemoglobin complex, an Fe^{2+} ion has a co-ordination number of six.

A multidentate haem ligand and an oxygen ligand form co-ordinate bonds to Fe^{2+} enabling oxygen to be transported in the blood.

Define multidentate ligand.

[1 mark]

____

____

0 6 . 3 The concentration of an Fe^{2+} complex in blood can be measured by colorimetry using light of wavelength 830 nm

Calculate the energy change (ΔE), in J, when a d electron is excited in this Fe^{2+} complex.

Speed of light, $c = 3.00 \times 10^{8}$ m s^{-1}

Planck constant, $h = 6.63 \times 10^{-34}$ J s

[2 marks]

ΔE ______________________ J

Adding an aqueous solution of $EDTA^{4-}$ to blood removes $Ca^{2+}(aq)$

$$[Ca(H_2O)_6]^{2+} + EDTA^{4-} \rightarrow [Ca(EDTA)]^{2-} + 6H_2O$$

0 6 . 4 Explain, in terms of entropy and Gibbs free-energy, why the formation of the metal–EDTA complex in this reaction is **not** reversible.

[3 marks]

0 6 . 5 Describe a test-tube reaction to show that $EDTA^{4-}$ has removed calcium ions from an aqueous sample of calcium chloride with a concentration of 0.100 mol dm^{-3}

State the result of the test.

[2 marks]

Test

Result

9

Do not write outside the box

0 7 This question is about transition metals.

0 7 . 1 State why the d block metal zinc is **not** a transition metal.

[1 mark]

In the Contact process, vanadium(V) oxide (V_2O_5) acts as a heterogeneous catalyst in the oxidation of sulfur dioxide to sulfur trioxide.

0 7 . 2 Write **two** equations to show how V_2O_5 acts as a catalyst in the Contact process.

Explain why the catalyst is heterogeneous.

[3 marks]

Equation 1

Equation 2

Explanation

0 7 . 3 Suggest why it is important that any impurities are removed from the gas before sulfur dioxide is used in the Contact process.

[1 mark]

Do not write outside the box

Table 5 shows the colours of aqueous solutions of vanadium ions.

Table 5

Aqueous vanadium ion	$VO_2^+(aq)$	$VO^{2+}(aq)$	$V^{3+}(aq)$	$V^{2+}(aq)$
Colour	Yellow	Blue	Green	Purple

Table 6 shows the standard electrode potentials of aqueous vanadium ions.

Table 6

Electrode half-equation	$E^\ominus$ / V
$VO_2^+(aq) + 2H^+(aq) + e^- \rightarrow VO^{2+}(aq) + H_2O(l)$	+1.00
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$	+0.34
$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$	–0.26
$V^{2+}(aq) + 2e^- \rightarrow V(s)$	–1.20

0 7 . 4 An excess of powdered zinc is added to an aqueous solution containing VO_2^+ ions. This causes a series of colour changes.

When zinc stops reacting, the solution is purple.

Use **Table 5** and **Table 6** to suggest a value for the standard electrode potential for the reduction of Zn^{2+}

Write a half-equation, with state symbols, for the reduction of Zn^{2+}

Write an overall equation for the reaction between zinc and VO_2^+ ions to form the purple solution.

[3 marks]

Electrode potential for reduction of Zn^{2+} ______________________________ V

Half equation

____

Overall equation

____

Turn over ►

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0 7 . 5 Transition metal complexes can have a range of shapes.

Table 7 shows the different shapes shown by transition metal complexes.

Tick (✓) one box in each row of **Table 7** to show the shape of each transition metal complex.

[2 marks]

Table 7

Complex	**Linear**	**Octahedral**	**Tetrahedral**	**Square Planar**
$FeCl_4^-$				
$Pt(NH_3)_2Cl_2$				
$[Ag(NH_3)_2]^+$				
$[Cu(NH_2CH_2CH_2NH_2)_2(H_2O)_2]^{2+}$				

___
10

Turn over for the next question

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0 8 This question is about aluminium.

0 8 . 1 Which block of the Periodic Table contains aluminium?

Tick (✓) **one** box.

[1 mark]

s block ☐

p block ☐

d block ☐

f block ☐

0 8 . 2 Give an equation for the reaction of aluminium with oxygen.

[1 mark]

____

0 8 . 3 Aluminium hydroxide, $Al(H_2O)_3(OH)_3$, reacts with acids and with alkalis.

State the word that describes this property.

Give an equation to show aluminium hydroxide reacting with an acid.

Give an equation to show aluminium hydroxide reacting with an alkali.

[3 marks]

Word to describe the property ______________________________

Equation for aluminium hydroxide reacting with an acid

____

Equation for aluminium hydroxide reacting with an alkali

____

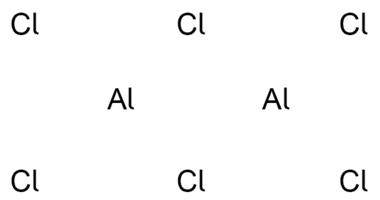
Do not write outside the box

0 8 . 4 **Figure 5** shows the atoms in a molecule of Al_2Cl_6

Complete **Figure 5**.
Use — to show covalent bonds and → to show co-ordinate bonds.

[1 mark]

Figure 5



0 8 . 5 Explain why $[Al(H_2O)_6]^{3+}$(aq) forms an acidic solution in water.
Include an equation in your answer.

[3 marks]

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0 8 . 6 Aqueous sodium carbonate is added to aqueous aluminium sulfate in a test tube.

Describe what is observed in this reaction.

Write an equation.

[3 marks]

Observation(s)

____

____

Equation

____

12

0 9 This question is about a redox titration.

Iodate ions (IO_3^-) react with iodide ions (I^-) in acidic conditions to produce iodine (I_2).

$$IO_3^- + 6H^+ + 5I^- \rightarrow 3I_2 + 3H_2O$$

Sodium thiosulfate ($Na_2S_2O_3$) reacts with iodine.

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$

- 0.545 g of KIO_3 is dissolved in water and made up to a 250 cm^3 solution in a volumetric flask.
- A 25.0 cm^3 portion of the solution is transferred to a conical flask.
- An excess of sulfuric acid is added to the conical flask.
- An excess of potassium iodide solution is added to the conical flask. The solution turns dark brown because iodine is produced.
- Sodium thiosulfate is added from a burette until the solution turns colourless. This shows that all the iodine has reacted.

Table 8 shows the titration results.

Table 8

	1	**2**	**3**
Final volume / cm^3	27.00	47.85	33.10
Initial volume / cm^3	3.65	25.35	10.55
Titre / cm^3	23.35	22.50	22.55

0 9 . 1 Use suitable values from **Table 8** to calculate the mean titre.

[1 mark]

Mean titre ______________________________ cm^3

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0 9 . 2 Use these equations to deduce how many moles of $S_2O_3^{2-}$ react with the iodine produced from the reaction of one mole of IO_3^-

$$IO_3^- + 6H^+ + 5I^- \rightarrow 3I_2 + 3H_2O$$

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$

[1 mark]

Moles of $S_2O_3^{2-}$ required ____________________

0 9 . 3 Use your answers to Question **09.1** and Question **09.2** to calculate the concentration, in mol dm^{-3}, of the sodium thiosulfate solution.

(If you were unable to answer Question **09.2** then you should assume that the moles of $S_2O_3^{2-}$ required is 2.5
This is **not** the correct answer.)

[4 marks]

Concentration ______________ mol dm^{-3}

Question 9 continues on the next page

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The method from page 22 is repeated below.

Iodate ions (IO_3^-) react with iodide ions (I^-) in acidic conditions to produce iodine (I_2).

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- An excess of potassium iodide solution is added to the conical flask. The solution turns dark brown because iodine is produced.
- Sodium thiosulfate is added from a burette until the solution turns colourless. This shows that all the iodine has reacted.

0 9 . 4 Explain why there should be an excess of potassium iodide and sulfuric acid in the conical flask.

[1 mark]

0 9 . 5 Explain why the exact amount of potassium iodide and sulfuric acid used in the conical flask does **not** need to be known.

[1 mark]

8

END OF QUESTIONS

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