

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.

A-level CHEMISTRY

Paper 1 Inorganic and Physical Chemistry

Time allowed: 2 hours

Materials

For this paper you must have:

- the Periodic Table/Data Booklet, provided as an insert (enclosed)
- 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 105.

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



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

0 1

This question is about enthalpy changes for calcium chloride and magnesium chloride.

0 1 . 1

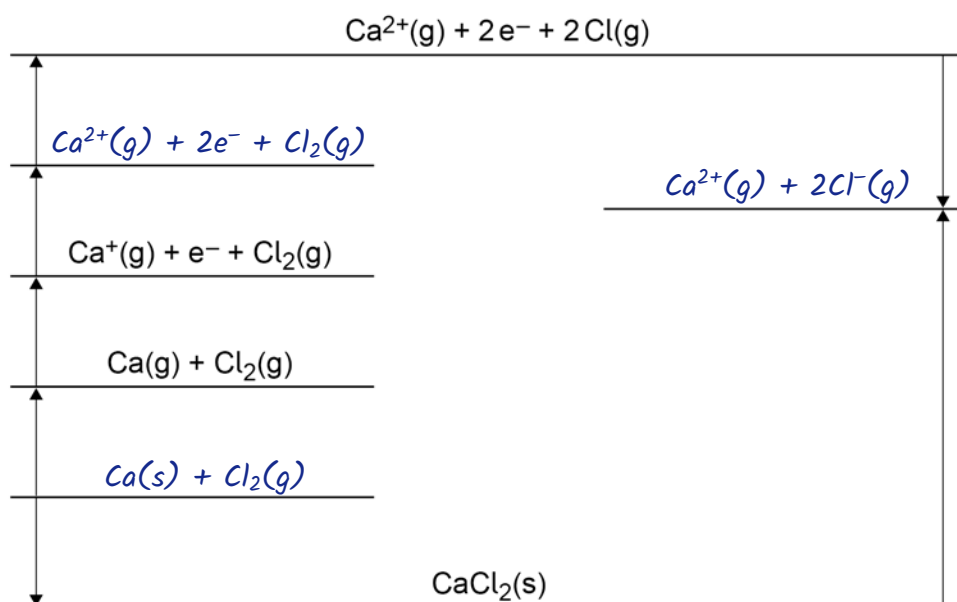
State the meaning of the term enthalpy change.

[1 mark]

The heat energy change (enthalpy change) for a reaction measured at constant pressure.

Figure 1 shows an incomplete Born–Haber cycle for the formation of calcium chloride.

Figure 1



0 1 . 2

Complete **Figure 1** by writing the formulas, including state symbols, of the appropriate species on each of the three blank lines.

[3 marks]



0 1 3

Table 1 shows some enthalpy data.

Table 1

	Enthalpy change / kJ mol^{-1}
Enthalpy of formation of calcium chloride	-795
Enthalpy of atomisation of calcium	+193
First ionisation energy of calcium	+590
Second ionisation energy of calcium	+1150
Enthalpy of atomisation of chlorine	+121
Electron affinity of chlorine	-364

Use **Figure 1** and the data in **Table 1** to calculate a value for the enthalpy of lattice dissociation of calcium chloride.

[2 marks]

$$\Delta H_f = \Delta H_{at}(\text{Ca}) + IE_1 + IE_2 + 2\Delta H_{at}(\text{Cl}) + 2EA(\text{Cl}) - \Delta H_{LD}$$

$$-795 = 193 + 590 + 1150 + (2 \times 121) + (2 \times -364) - \Delta H_{LD}$$

$$-795 = 1447 - \Delta H_{LD}$$

$$\Delta H_{LD} = 1447 + 795 = +2242 \text{ kJ mol}^{-1}$$

Enthalpy of lattice dissociation +2242 kJ mol^{-1}

Question 1 continues on the next page

Turn over ►



0 1 . 4

Magnesium chloride dissolves in water.

Give an equation, including state symbols, to represent the process that occurs when the enthalpy of solution of magnesium chloride is measured.

[1 mark]



0 1 . 5

Table 2 shows some enthalpy data.

Table 2

	Enthalpy change / kJ mol^{-1}
Enthalpy of lattice dissociation of MgCl_2	+2493
Enthalpy of hydration of $\text{Mg}^{2+}(\text{g})$	−1920
Enthalpy of hydration of $\text{Cl}^-(\text{g})$	−364

Use your answer to Question 01.4 and the data in Table 2 to calculate a value for the enthalpy of solution of magnesium chloride.

[2 marks]

$$\begin{aligned}\Delta H_{\text{soln}}(\text{MgCl}_2) &= \Delta H_{\text{latt diss}} + \Delta H_{\text{hyd}}(\text{Mg}^{2+}) + 2\Delta H_{\text{hyd}}(\text{Cl}^-) \\ &= 2493 + (-1920) + (2 \times -364) \\ &= 2493 - 1920 - 728 = -155 \text{ kJ mol}^{-1}\end{aligned}$$

Enthalpy of solution −155 kJ mol^{-1}

0 1 . 6

The enthalpy of hydration of $\text{Ca}^{2+}(\text{g})$ is $-1650 \text{ kJ mol}^{-1}$

Suggest why this value is less exothermic than that of $\text{Mg}^{2+}(\text{g})$

[2 marks]

The Ca^{2+} ion is bigger than the Mg^{2+} ion (same charge), so it has a lower charge to size ratio (lower charge density).

So there is a weaker attraction between Ca^{2+} and the δ^- oxygen atoms of the water molecules.



0 2

This question is about atomic structure.

0 2 . 1

Define the mass number of an atom.

[1 mark]

The number of protons plus neutrons in the nucleus of the atom.

0 2 . 2

Complete **Table 3** to show the numbers of neutrons and electrons in the species shown.

[2 marks]

Table 3

	Number of protons	Number of neutrons	Number of electrons
^{46}Ti	22	24	22
$^{49}\text{Ti}^{2+}$	22	27	20

0 2 . 3

A sample of titanium contains four isotopes, ^{46}Ti , ^{47}Ti , ^{48}Ti and ^{49}Ti

This sample has a relative atomic mass of 47.8

In this sample the ratio of abundance of isotopes ^{46}Ti , ^{47}Ti and ^{49}Ti is 2:2:1

Calculate the percentage abundance of ^{46}Ti in this sample.

[3 marks]

Let abundance of $^{49}\text{Ti} = y$, so $^{46}\text{Ti} = 2y$, $^{47}\text{Ti} = 2y$, $^{48}\text{Ti} = 100 - 5y$

$$47.8 = \frac{(46 \times 2y) + (47 \times 2y) + (48 \times (100 - 5y)) + (49 \times y)}{100}$$

$$4780 = 92y + 94y + 4800 - 240y + 49y$$

$$4780 = 4800 - 5y, \text{ so } 5y = 20, y = 4$$

$$\text{abundance of } ^{46}\text{Ti} = 2y = 2 \times 4 = 8 \%$$

Abundance of ^{46}Ti 8 %

6

Turn over ►



0 3

This question is about elements in Period 3 and their compounds.

0 3 . 1

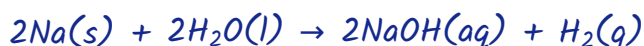
When a piece of sodium is added to 200 cm³ of water in a large beaker a vigorous reaction occurs. The temperature of the water increases by 25 °C

Give an equation, including state symbols, for the reaction of sodium with water.

Suggest why it is dangerous to react a similar piece of sodium with 10 cm³ of water in a boiling tube.

[2 marks]

Equation



Why it is dangerous With only 10 cm³ of water the temperature will go up much more, so the reactants could shoot out of the tube (the hydrogen could ignite or the glass could shatter).

0 3 . 2

Give an equation for the reaction of phosphorus(V) oxide with water.

Suggest a pH for the solution formed.

[2 marks]

Equation



pH 0

0 3 . 3

Explain, in terms of crystal structure and bonding, why silicon(IV) oxide has a higher melting point than phosphorus(V) oxide.

[4 marks]

SiO₂ is macromolecular (giant covalent): every atom is joined to others by strong covalent bonds throughout the whole structure.

To melt SiO₂ these strong covalent bonds between atoms must be broken, which needs a lot of energy.

P₄O₁₀ is molecular (simple covalent molecules). Only the weak van der Waals forces between molecules have to be overcome to melt it, and



these break easily, so less energy is needed and its melting point is lower.

0 3 . 4

An element in Period 3 forms an oxide that is insoluble in water. This oxide reacts with sulfuric acid and with aqueous potassium hydroxide.

Give the formula for this oxide.

Give an equation for the reaction of this oxide with sulfuric acid.

[2 marks]

Formula *Al₂O₃*

Equation



0 3 . 5

Give the formula of a hydroxide of an element in Period 3 used in medicine.

[1 mark]



0 3 . 6

Identify the element in Period 3, from sodium to chlorine, that has the largest atomic radius.

[1 mark]

Sodium (Na)

12

Turn over for the next question

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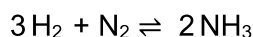


0 4

This question is about iron and its ions.

0 4 . 1

Discuss the role of iron as a heterogeneous catalyst in the Haber process.



Your answer should include:

- the meaning of the term heterogeneous catalyst
- how iron acts as a heterogeneous catalyst
- the factors that affect the efficiency and lifetime of the catalyst.

[6 marks]

Meaning: a heterogeneous catalyst is in a different phase (state) from the reactants. Here the iron is a solid while the hydrogen and nitrogen are gases. A catalyst speeds up the reaction without being used up (it is left chemically unchanged), because it provides an alternative route with a lower activation energy.

How iron acts: the reactants, hydrogen and nitrogen, are adsorbed onto the surface of the iron at the active sites. This weakens the bonds in the reactant molecules (especially the $\text{N}\equiv\text{N}$ triple bond), so the reaction takes place on the surface. The product, ammonia, then desorbs (leaves) from the surface of the iron, which frees the active sites so that more reactant molecules can be adsorbed.

Efficiency: because the reaction happens on the surface, the iron is used with a large surface area, for example as a powder or small pellets, or spread on a support medium (mesh). This gives more active sites and so a faster reaction for less iron.

Lifetime: the catalyst can be poisoned by impurities in the gases, for example sulfur (compounds). The sulfur is strongly adsorbed onto (binds to) the iron surface and blocks the active sites, so the reactants can no longer be adsorbed. This reduces the efficiency of the catalyst and



shortens its lifetime, so the gases are purified first.

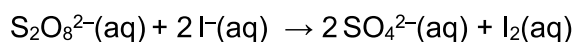
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0 4 . 2

Fe^{2+} ions catalyse the reaction between peroxodisulfate(VI) ions and iodide ions in aqueous solution.



Explain why this reaction is slow before the catalyst is added.
Give **two** equations to show how Fe^{2+} ions catalyse this reaction.

[4 marks]

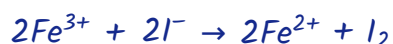
Why reaction is slow before catalyst added Both reactant ions ($\text{S}_2\text{O}_8^{2-}$ and I^-) are negative, so the two negative ions repel each other.

So the activation energy is high and few collisions are successful.

Equation 1



Equation 2



0 4 . 3

Give a reason why Zn^{2+} ions do **not** catalyse the reaction in Question **04.2**.

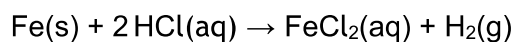
[1 mark]

Zinc ions have only one oxidation state (Zn^{2+} is the only ion), so they cannot be oxidised to Zn^{3+} .



0 4 . 4

Iron reacts with dilute hydrochloric acid to form iron(II) chloride and hydrogen.



A 0.998 g sample of pure iron is added to 30.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid.

One of these reagents is in excess and the other reagent limits the amount of hydrogen produced in the reaction.

Calculate the maximum volume, in m³, of hydrogen gas produced at 30 °C and 100 kPa.

Give your answer to 3 significant figures.

In your answer you should identify the limiting reagent in the reaction.

The gas constant, $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$

[6 marks]

$$n(\text{Fe}) = 0.998 \div 55.8 = 0.0179 \text{ mol}$$

$$n(\text{HCl}) = 30.0 \times 10^{-3} \times 1.00 = 0.0300 \text{ mol}$$

$$0.0179 \text{ mol Fe needs } 2 \times 0.0179 = 0.0358 \text{ mol HCl, but only } 0.0300 \text{ mol}$$

so HCl is the limiting reagent (iron is in excess)

$$n(\text{H}_2) = 0.0300 \div 2 = 0.0150 \text{ mol}$$

$$T = 30 + 273 = 303 \text{ K, } p = 100 \text{ kPa} = 100\,000 \text{ Pa}$$

$$pV = nRT, \text{ so } V = nRT \div p$$

$$V = (0.0150 \times 8.31 \times 303) \div 100\,000$$

$$= 3.78 \times 10^{-4} \text{ m}^3 \text{ (3 s.f.)}$$

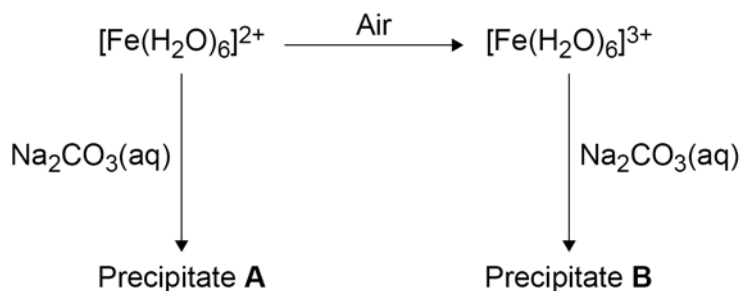
Volume of hydrogen 3.78 × 10⁻⁴ m³

Turn over ►



Figure 2 shows some reactions of iron ions in aqueous solution.

Figure 2



0 4 . 5 Identify **A** and state its colour.

[2 marks]

Identity FeCO_3 (iron(II) carbonate)

Colour Green

0 4 . 6 Give the formula of **B** and state its colour.

Give an ionic equation for the reaction of $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$ with aqueous Na_2CO_3 to form **B**.

[3 marks]

Formula $\text{Fe}(\text{H}_2\text{O})_3(\text{OH})_3$

Colour Brown

Ionic equation



0 4 . 7

Explain why an aqueous solution containing $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$ ions has a lower pH than an aqueous solution containing $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ ions.

[3 marks]

Fe^{3+} is smaller than Fe^{2+} and has a greater charge, so Fe^{3+} has a greater charge to size ratio (charge density).

So Fe^{3+} ions are more polarising: they polarise (attract electrons from) the water ligands more strongly.

So the O-H bonds in the water ligands are weakened and break more in the Fe^{3+} solution, and more H^+ ions are released, giving a higher $[\text{H}^+]$ and a lower pH.

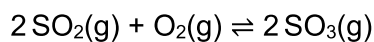
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0 5

This question is about the equilibrium



0 5 . 1

State and explain the effect, if any, of a decrease in overall pressure on the equilibrium yield of SO_3

[3 marks]

Effect The equilibrium yield of SO_3 decreases.

Explanation There are more moles of gas on the left (3) than on the right (2). So the equilibrium shifts to the left, to the side with more gas molecules, to oppose the decrease in pressure.

0 5 . 2

A 0.460 mol sample of SO_2 is mixed with a 0.250 mol sample of O_2 in a sealed container at a constant temperature. When equilibrium is reached at a pressure of 215 kPa, the mixture contains 0.180 mol of SO_3

Calculate the partial pressure, in kPa, of SO_2 in this equilibrium mixture.

[4 marks]

at equilibrium:

$$n(\text{SO}_2) = 0.460 - 0.180 = 0.280 \text{ mol}$$

$$n(\text{O}_2) = 0.250 - (0.180 \div 2) = 0.250 - 0.090 = 0.160 \text{ mol}$$

$$n(\text{SO}_3) = 0.180 \text{ mol}$$

$$\text{total moles} = 0.280 + 0.160 + 0.180 = 0.620 \text{ mol}$$

$$\text{mole fraction of } \text{SO}_2 = 0.280 \div 0.620 = 0.4516$$

$$pp(\text{SO}_2) = \text{mole fraction} \times \text{total pressure}$$

$$= 0.4516 \times 215 = 97.1 \text{ kPa}$$

Partial pressure of SO_2 97.1 kPa



Question 5 continues on the next page

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0 5 3

A different mixture of SO_2 and O_2 reaches equilibrium at a different temperature.

Table 4 shows the partial pressures of the gases at equilibrium.

Table 4

Gas	Partial pressure / kPa
SO_2	1.67×10^2
O_2	1.02×10^2
SO_3	1.85×10^2

Give an expression for the equilibrium constant (K_p) for this reaction.

Calculate the value of the equilibrium constant for this reaction and give its units.

[3 marks]

$$K_p = \frac{(p_{\text{SO}_3})^2}{(p_{\text{SO}_2})^2 \times (p_{\text{O}_2})}$$

$$K_p = \frac{(1.85 \times 10^2)^2}{(1.67 \times 10^2)^2 \times (1.02 \times 10^2)}$$

$$= 34\,225 \div 2\,844\,678 = 0.01203$$

$$\text{units: } \text{kPa}^2 \div (\text{kPa}^2 \times \text{kPa}) = \text{kPa}^{-1}$$

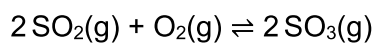
$$K_p \quad 1.20 \times 10^{-2}$$

$$\text{Units} \quad \text{kPa}^{-1}$$



0 5 . 4

What is the effect on the value of K_p if the pressure of this equilibrium mixture is increased at a constant temperature?

**[1 mark]**

Tick (✓) **one** box.

The value of K_p

increases.

☐

stays the same.

☒

decreases.

☐

11

Turn over for the next question

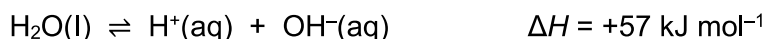
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0 6

This question is about pH.

Pure water dissociates slightly.



The equilibrium constant, $K_c = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$

The ionic product of water, $K_w = [\text{H}^+][\text{OH}^-]$

0 6

1

Explain why $[\text{H}_2\text{O}]$ is not shown in the K_w expression.

[1 mark]

Because $[\text{H}_2\text{O}]$ is (almost) constant: it is very large compared with $[\text{H}^+]$ and $[\text{OH}^-]$, so it is incorporated into K_w ($K_w = K_c[\text{H}_2\text{O}]$).

Table 5 shows how K_w varies with temperature.

Table 5

Temperature / °C	$K_w / \text{mol}^2 \text{ dm}^{-6}$
10	2.93×10^{-15}
20	6.81×10^{-15}
25	1.00×10^{-14}
30	1.47×10^{-14}
50	5.48×10^{-14}

0 6

2

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

[2 marks]

The forward reaction is endothermic ($\Delta H = +57 \text{ kJ mol}^{-1}$). So as the temperature increases the equilibrium shifts to the right-hand side to oppose the temperature increase, so $[\text{H}^+]$ and $[\text{OH}^-]$ increase and K_w increases.



0 6 . 3

Give the expression for pH.

Calculate the pH of pure water at 50 °C

Give your answer to 2 decimal places.

Explain why water is neutral at 50 °C

[4 marks]Expression $pH = -\log_{10}[H^+]$

Calculation

$$\text{In pure water } [H^+] = [OH^-], \text{ so } K_w = [H^+]^2$$

$$[H^+]^2 = 5.48 \times 10^{-14}$$

$$[H^+] = \sqrt{(5.48 \times 10^{-14})} = 2.34 \times 10^{-7} \text{ mol dm}^{-3}$$

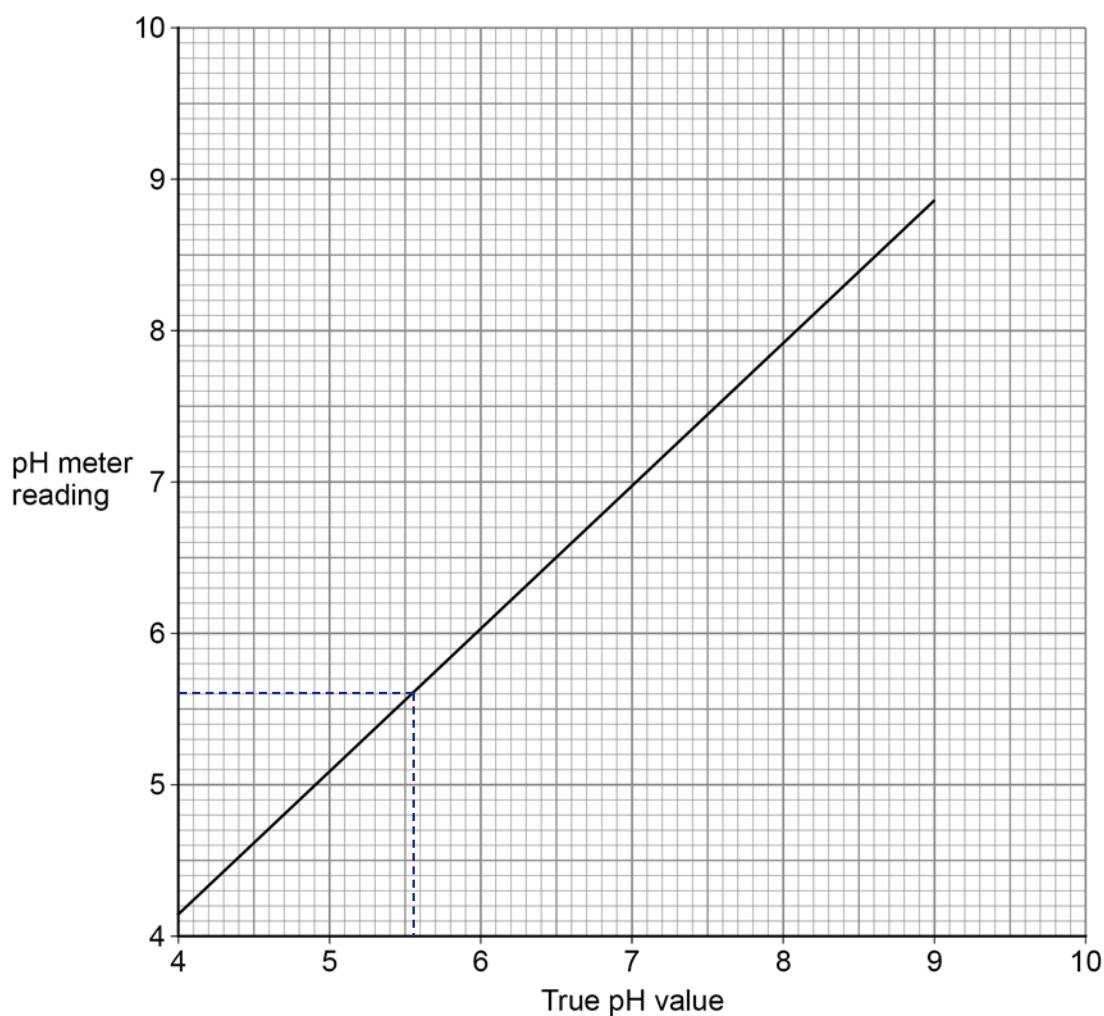
$$pH = -\log_{10}(2.34 \times 10^{-7}) = 6.63$$

pH 6.63Explanation $[H^+] = [OH^-]$, because the dissociation of each water molecule gives one H^+ and one OH^- .**Question 6 continues on the next page****Turn over ►**

A pH meter is calibrated using a calibration graph.
To create the calibration, the pH meter is used to measure the pH of separate solutions, each with a known, accurate pH.

Figure 3 shows the calibration graph.

Figure 3



0 6 . 4 Use **Figure 3** to give the true pH value when the pH meter reading is 5.6

[1 mark]

5.55

0 6 . 5 Suggest why the pH probe is washed with distilled water between each of the calibration measurements.

[1 mark]

To wash off any residual solution so the different solutions do not contaminate each other and affect the reading.



0 6 . 6

The calibrated pH meter is used to monitor the pH during a titration of hydrochloric acid with sodium hydroxide.

Explain why the volume of sodium hydroxide solution added between each pH measurement is smaller as the end point of the titration is approached.

[1 mark]

Near the end point there is a very large change in pH for a small volume added, so small additions avoid missing the end point.

Figure 4 shows the pH curve for a titration of hydrochloric acid with sodium hydroxide solution.

Figure 4

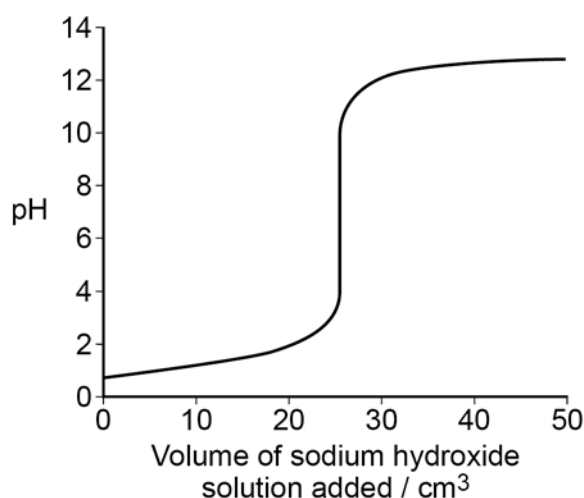


Table 6 shows data about some indicators.

Table 6

Indicator	pH range	Colour at low pH	Colour at high pH
Bromocresol green	3.8 – 5.4	yellow	blue
Phenol red	6.8 – 8.4	yellow	red
Thymolphthalein	9.3 – 10.5	colourless	blue

The student plans to do the titration again using one of the indicators in **Table 6** to determine the end point.

0 6 . 7

State why all three of the indicators in **Table 6** are suitable for this titration.

[1 mark]

Each indicator changes colour (pH range) within the steep (vertical) part of the titration curve, between pH 3 and 11.

Turn over ►



0 6 . 8

36.25 cm³ of 0.200 mol dm⁻³ sodium hydroxide solution are added to 25.00 cm³ of 0.150 mol dm⁻³ hydrochloric acid.

Calculate the pH of the final solution at 25 °C

$K_w = 1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ at 25 °C

[5 marks]

$$n(\text{OH}^-) = 36.25 \times 0.200 \div 1000 = 7.25 \times 10^{-3} \text{ mol}$$

$$n(\text{H}^+) = 25.00 \times 0.150 \div 1000 = 3.75 \times 10^{-3} \text{ mol}$$

$$\text{excess } n(\text{OH}^-) = 7.25 \times 10^{-3} - 3.75 \times 10^{-3} = 3.50 \times 10^{-3} \text{ mol}$$

$$\text{total volume} = 36.25 + 25.00 = 61.25 \text{ cm}^3$$

$$[\text{OH}^-] = 3.50 \times 10^{-3} \div (61.25 \times 10^{-3}) = 5.71 \times 10^{-2} \text{ mol dm}^{-3}$$

$$[\text{H}^+] = K_w \div [\text{OH}^-] = 1.00 \times 10^{-14} \div 5.71 \times 10^{-2}$$

$$= 1.75 \times 10^{-13} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log_{10}(1.75 \times 10^{-13}) = 12.76$$

pH 12.76

16



Turn over for the next question

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0 7

This question is about thermodynamics.
Consider the reaction shown.

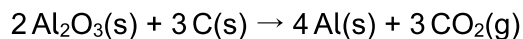


Table 7 shows some thermodynamic data.

Table 7

Substance	$\text{Al}_2\text{O}_3(\text{s})$	$\text{Al}(\text{s})$	$\text{C}(\text{s})$	$\text{CO}_2(\text{g})$
$\Delta_f H^\ominus / \text{kJ mol}^{-1}$	-1669	0	0	-394
$S^\ominus / \text{J K}^{-1} \text{mol}^{-1}$	51	28	6	214

0 7 . 1

Explain why the standard entropy value for carbon dioxide is greater than that for carbon.

[1 mark]

CO₂ is a gas, so its particles are much more disordered (more random, more ways of arranging energy) than in solid carbon.

0 7 . 2

State the temperature at which the standard entropy of aluminium is $0 \text{ J K}^{-1} \text{mol}^{-1}$

[1 mark]

0 K (absolute zero)



0 7 3

Use the equation and the data in **Table 7** to calculate the minimum temperature, in K, at which this reaction becomes feasible.

[7 marks]

$$\Delta H = \Sigma \Delta_f H(\text{products}) - \Sigma \Delta_f H(\text{reactants})$$

$$\Delta H = (3 \times -394) - (2 \times -1669) = -1182 + 3338 = +2156 \text{ kJ mol}^{-1}$$

$$\Delta S = \Sigma S(\text{products}) - \Sigma S(\text{reactants})$$

$$\Delta S = (4 \times 28 + 3 \times 214) - (2 \times 51 + 3 \times 6)$$

$$= 754 - 120 = +634 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$= 0.634 \text{ kJ K}^{-1} \text{ mol}^{-1}$$

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

reaction becomes feasible when $\Delta G = 0$, so $\Delta H = T\Delta S$

$$T = \Delta H \div \Delta S$$

$$T = 2156 \div 0.634 = 3401 \text{ K}$$

Minimum temperature 3401 K

9

Turn over ►



0 8

This question is about electrode potentials and electrochemical cells.

0 8 . 1

State the meaning of the term electrochemical series.

[1 mark]

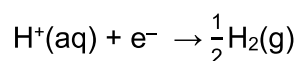
A list of electrode potentials ($E^\ominus$ values) arranged in numerical order.

Table 8 shows some electrode potentials.

Table 8

	$E^\ominus / \text{V}$
$[\text{Fe}(\text{H}_2\text{O})_6]^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s}) + 6\text{H}_2\text{O}(\text{l})$	-0.44
$\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \frac{1}{2}\text{H}_2(\text{g})$	0.00
$[\text{Co}(\text{NH}_3)_6]^{3+}(\text{aq}) + \text{e}^- \rightarrow [\text{Co}(\text{NH}_3)_6]^{2+}(\text{aq})$	+0.11
$[\text{Fe}(\text{H}_2\text{O})_6]^{3+}(\text{aq}) + \text{e}^- \rightarrow [\text{Fe}(\text{H}_2\text{O})_6]^{2+}(\text{aq})$	+0.77
$\text{VO}_2^+(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+1.00
$[\text{Co}(\text{H}_2\text{O})_6]^{3+}(\text{aq}) + \text{e}^- \rightarrow [\text{Co}(\text{H}_2\text{O})_6]^{2+}(\text{aq})$	+1.81

0 8 . 2

State **two** conditions needed for the following half-cell to have $E^\ominus = 0.00 \text{ V}$ 

[1 mark]

Temperature 298 K (25 °C); $[\text{H}^+] = 1.00 \text{ mol dm}^{-3}$ (and H_2 at 100 kPa).

0 8 . 3

Identify the weakest reducing agent in **Table 8**.

[1 mark]

$[\text{Co}(\text{H}_2\text{O})_6]^{2+}$



0 8 . 4

Use half-equations from **Table 8** to deduce an equation for the reduction of VO_2^+ to form VO^{2+} in aqueous solution by iron.

[2 marks]



0 8 . 5

Use data from **Table 8** to explain why $[\text{Co}(\text{H}_2\text{O})_6]^{3+}(\text{aq})$ will undergo a redox reaction with $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}(\text{aq})$

Give an equation for this reaction.

[2 marks]

Explanation $E^\ominus$ for $\text{Co}^{3+}/\text{Co}^{2+}$ (+1.81 V) is more positive than $E^\ominus$ for $\text{Fe}^{3+}/\text{Fe}^{2+}$ (+0.77 V), so $[\text{Co}(\text{H}_2\text{O})_6]^{3+}$ oxidises $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ (EMF = +1.04 V).

Equation



0 8 . 6

Suggest why the **two** cobalt(III) complex ions in **Table 8** have different electrode potentials.

[1 mark]

They have different ligands (NH_3 compared with H_2O).

Turn over for the next question

Turn over ►



09

This question is about the development of lithium cells.
The value of $E^\ominus$ for lithium suggests that a lithium cell could have a large EMF.

Table 9 shows some electrode potential data.

Table 9

	$E^\ominus / \text{V}$
$\text{Li}^+(\text{aq}) + \text{e}^- \rightarrow \text{Li}(\text{s})$	-3.04
$2 \text{H}_2\text{O}(\text{l}) + 2 \text{e}^- \rightarrow \text{H}_2(\text{g}) + 2 \text{OH}^-(\text{aq})$	-0.83
$\frac{1}{2} \text{I}_2(\text{s}) + \text{e}^- \rightarrow \text{I}^-(\text{aq})$	+0.54

09.1

Use data in **Table 9** to explain why an aqueous electrolyte is **not** used for a lithium cell.

[2 marks]

Lithium would react with the water in the electrolyte (Li reduces water to hydrogen), because $E^\ominus$ for Li^+/Li (-3.04 V) is more negative than $E^\ominus$ for $\text{H}_2\text{O}/\text{H}_2, \text{OH}^-$ (-0.83 V).

09.2

In the 1970s lithium-iodine cells became a common power source for heart pacemakers. Lithium iodide is the final product of the cell reaction.

Use the data in **Table 9** to calculate the cell EMF of a standard lithium-iodine cell.

[1 mark]

$$EMF = 0.54 - (-3.04) = 3.58 \text{ V}$$

09.3

An EMF value for a commercial lithium-iodine cell is 2.80 V

Suggest why this value is different from the value calculated in Question 09.2.

[1 mark]

The commercial cell does not operate under standard conditions (e.g. it is non-aqueous).



0 9 . 4

In some lithium cells, lithium perchlorate (LiClO_4) is used as the electrolyte.

Deduce the oxidation state of chlorine in LiClO_4

[1 mark]

$+7$

In other lithium cells, lithium cobalt oxide electrodes **and** lithium electrodes are used.

0 9 . 5

Give an equation for the reaction that occurs at the positive lithium cobalt oxide electrode.

[1 mark]

$\text{Li}^+ + \text{CoO}_2 + \text{e}^- \rightarrow \text{LiCoO}_2$

0 9 . 6

Give an equation for the reaction that occurs at the negative lithium electrode.

[1 mark]

$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$

7

END OF QUESTIONS



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