

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

Tuesday 2 June 2020

Afternoon

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	
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6	
7	
8	
9	
10	
11	
TOTAL	



J U N 2 0 7 4 0 5 1 0 1

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

0 1

This question is about enthalpy changes.

0 1 . 1

Figure 1 shows a Born–Haber cycle for the formation of strontium chloride, SrCl_2

Figure 1

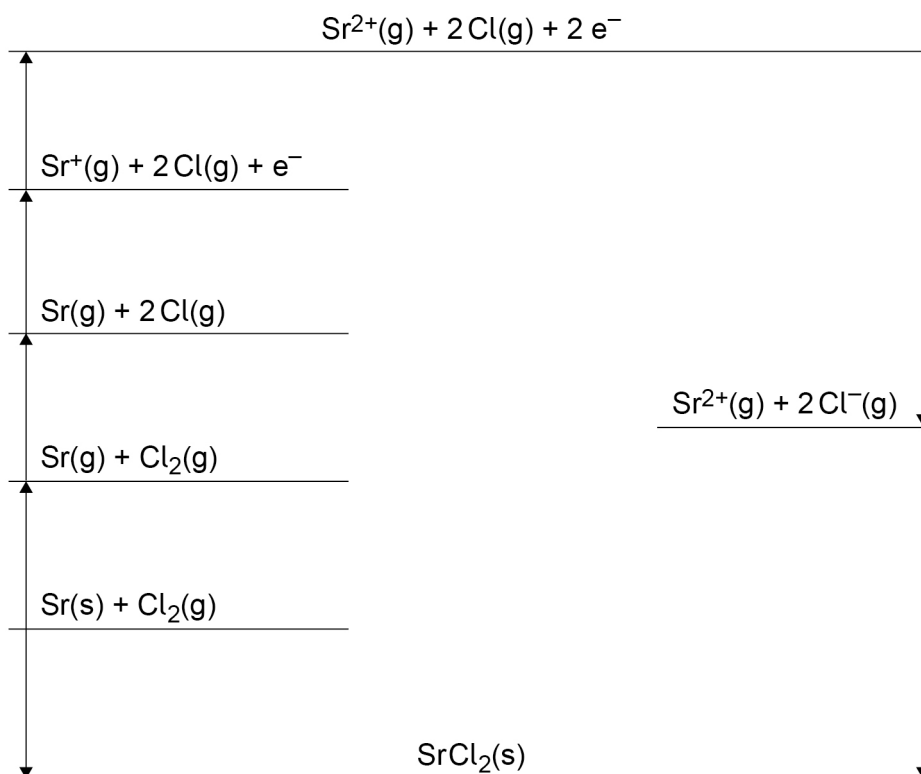


Table 1 shows some thermodynamic data.

Table 1

	Enthalpy change / kJ mol^{-1}
First ionisation energy of strontium	+548
Second ionisation energy of strontium	+1060
Enthalpy of atomisation of chlorine	+121
Enthalpy of atomisation of strontium	+164
Enthalpy of formation of strontium chloride	−828
Enthalpy of lattice formation of strontium chloride	−2112



Use the data in **Table 1** to calculate a value for the electron affinity of chlorine.

[3 marks]

$$\Delta_f H = \Delta_a H(\text{Sr}) + 2\Delta_a H(\text{Cl}) + 1E_1 + 1E_2 + 2EA(\text{Cl}) + \Delta_{LE} H$$

$$-828 = 164 + (2 \times 121) + 548 + 1060 + 2EA + (-2112)$$

$$2EA = -828 - 164 - 242 - 548 - 1060 + 2112 = -730$$

$$EA = -730 \div 2 = -365 \text{ kJ mol}^{-1}$$

Electron affinity -365 kJ mol⁻¹

0 1 2

Draw a line from **each** substance to the enthalpy of lattice formation of that substance.

[1 mark]

Substance	Enthalpy of lattice formation / kJ mol ⁻¹
MgCl ₂	-2018
MgO	-2493
BaCl ₂	-3889

Question 1 continues on the next page

Turn over ►



Table 2 shows the theoretical lattice enthalpy, based on a perfect ionic model, and an experimental value for the enthalpy of lattice formation of silver chloride.

Table 2

	Theoretical	Experimental
Enthalpy of lattice formation / kJ mol^{-1}	-770	-905

- 0 1 . 3** State why there is a difference between the theoretical and experimental values. **[1 mark]**

*Silver chloride has partial covalent character as well as ionic bonding:
the chloride ion is polarised (distorted) by the Ag^+ ion.*

- 0 1 . 4** **Table 3** shows enthalpy of hydration values for ions of some Group 1 elements.

Table 3

	$\text{Li}^+(\text{g})$	$\text{Na}^+(\text{g})$	$\text{K}^+(\text{g})$
Enthalpy of hydration / kJ mol^{-1}	-519	-406	-322

- Explain why the enthalpy of hydration becomes less exothermic from Li^+ to K^+ **[2 marks]**

From Li^+ to K^+ the ionic radius increases, so the charge density of the ion decreases.

So the electrostatic attraction between the positive metal ion and the lone pairs on the δ^- oxygen atoms of water molecules decreases.



0 1 5

Calcium bromide dissolves in water.

Table 4 shows some enthalpy data.**Table 4**

	Enthalpy change / kJ mol ⁻¹
Enthalpy of solution of calcium bromide	-110
Enthalpy of lattice formation of calcium bromide	-2176
Enthalpy of hydration of calcium ions	-1650

Use the data in **Table 4** to calculate the enthalpy of hydration, in kJ mol⁻¹, of bromide ions.

[3 marks]

$$\Delta_{sol}H = \Delta_{L(dissociation)}H + \Delta_{hyd}H(Ca^{2+}) + 2\Delta_{hyd}H(Br^{-})$$

$$-110 = (+2176) + (-1650) + 2\Delta_{hyd}H(Br^{-})$$

$$2\Delta_{hyd}H(Br^{-}) = -110 - 2176 + 1650 = -636$$

$$\Delta_{hyd}H(Br^{-}) = -636 \div 2 = -318 \text{ kJ mol}^{-1}$$

Enthalpy of hydration of bromide ions -318 kJ mol⁻¹

10

Turn over for the next question**Turn over ►**

0 2

This question is about the isotopes of chromium.

0 2 . 1

Give the meaning of the term relative atomic mass.

[2 marks]

*The average mass of one atom of an element**divided by 1/12 of the mass of one atom of carbon-12 (^{12}C).*

0 2 . 2

A sample of chromium containing the isotopes ^{50}Cr , ^{52}Cr and ^{53}Cr has a relative atomic mass of 52.1The sample contains 86.1% of the ^{52}Cr isotope.

Calculate the percentage abundance of each of the other two isotopes.

[4 marks]

$$\% ^{50}\text{Cr} + \% ^{53}\text{Cr} = 100 - 86.1 = 13.9\%$$

$$\text{Let } \% ^{53}\text{Cr} = x, \text{ so } \% ^{50}\text{Cr} = 13.9 - x$$

$$52.1 = [50(13.9 - x) + (52 \times 86.1) + 53x] \div 100$$

$$5210 = 695 - 50x + 4477.2 + 53x$$

$$3x = 37.8, \text{ so } x = \% ^{53}\text{Cr} = 12.6\%$$

$$\% ^{50}\text{Cr} = 13.9 - 12.6 = 1.3\%$$

Abundance of ^{50}Cr 1.3 %Abundance of ^{53}Cr 12.6 %

0 2 3

State, in terms of the numbers of fundamental particles, **one** similarity and **one** difference between atoms of ^{50}Cr and ^{53}Cr

[2 marks]

Similarity Same number of protons (24) and electrons

Difference Different number of neutrons: ^{50}Cr has 26 but ^{53}Cr has 29

The sample of chromium is analysed in a time of flight (TOF) mass spectrometer.

0 2 4

Give **two** reasons why it is necessary to ionise the isotopes of chromium before they can be analysed in a TOF mass spectrometer.

[2 marks]

1 So the ions can be accelerated by the electric field (towards the negative plate) so they all have the same kinetic energy.

2 So the ions create a current when they hit the detector, so they can be detected (size of current gives abundance).

Question 2 continues on the next page

Turn over ►



0 2 . 5

A $^{53}\text{Cr}^+$ ion travels along a flight tube of length 1.25 m
The ion has a constant kinetic energy (KE) of $1.102 \times 10^{-13} \text{ J}$

$$KE = \frac{mv^2}{2}$$

m = mass of the ion / kg

v = speed of ion / m s^{-1}

Calculate the time, in s, for the $^{53}\text{Cr}^+$ ion to travel down the flight tube to reach the detector.

The Avogadro constant, $L = 6.022 \times 10^{23} \text{ mol}^{-1}$

[5 marks]

$$m = 53 \div (6.022 \times 10^{23}) = 8.80 \times 10^{-23} \text{ g} = 8.80 \times 10^{-26} \text{ kg}$$

$$KE = \frac{1}{2}mv^2, \text{ so } v^2 = 2KE \div m$$

$$v^2 = (2 \times 1.102 \times 10^{-13}) \div (8.80 \times 10^{-26}) = 2.504 \times 10^{12}$$

$$v = \sqrt{(2.504 \times 10^{12})} = 1.58 \times 10^6 \text{ m s}^{-1}$$

$$v = d \div t, \text{ so } t = d \div v = 1.25 \div (1.58 \times 10^6)$$

$$t = 7.90 \times 10^{-7} \text{ s}$$

Time 7.90×10^{-7} s

15



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ANSWER IN THE SPACES PROVIDED**

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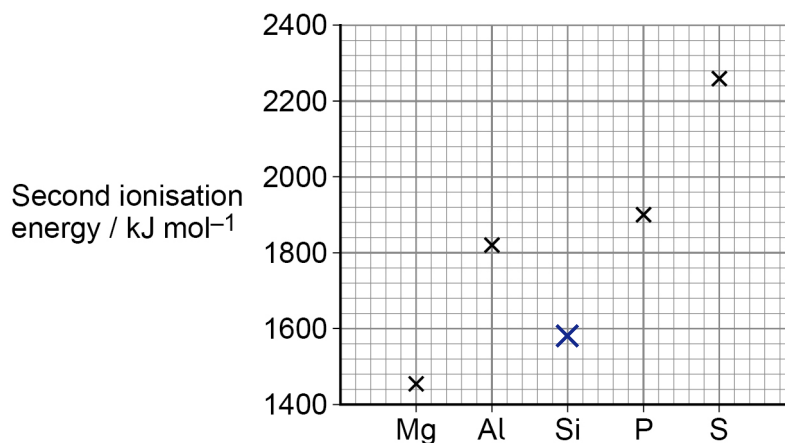


0 3

This question is about Period 3 elements.

Figure 2 shows the **second** ionisation energies of some elements in Period 3.

Figure 2



0 3 . 1

Draw a cross (x) on **Figure 2** to show the **second** ionisation energy of silicon.

[1 mark]

0 3 . 2

Identify the element in Period 3, from sodium to argon, that has the highest **second** ionisation energy.

Give an equation, including state symbols, to show the process that occurs when the **second** ionisation energy of this element is measured.

If you were unable to identify the element you may use the symbol **Q** in your equation.

[2 marks]

Element Sodium (Na)

Equation



0 3 . 3

Explain why the atomic radius decreases across Period 3, from sodium to chlorine.

[2 marks]

Across the period the number of protons increases, so the nuclear charge increases.

Electrons are added to the same (outer) shell, so shielding is similar, so the outer electrons are pulled closer to the nucleus.



03.4

Identify the element in Period 3, from sodium to chlorine, that has the highest electronegativity.

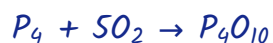
[1 mark]

Chlorine (Cl)

03.5

Phosphorus burns in air to form phosphorus(V) oxide.
Give an equation for this reaction.

[1 mark]



7

Turn over for the next question

Turn over ►



0 4

Propanoic acid ($\text{C}_2\text{H}_5\text{COOH}$) is a weak acid.

The acid dissociation constant (K_a) for propanoic acid is $1.35 \times 10^{-5} \text{ mol dm}^{-3}$ at 25°C

0 4 . 1

State the meaning of the term weak acid.

[1 mark]

An acid that only partially (slightly) dissociates into H^+ ions in water.

0 4 . 2

Give an expression for the acid dissociation constant for propanoic acid.

[1 mark]

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_5\text{COO}^-]}{[\text{C}_2\text{H}_5\text{COOH}]}$$

0 4 . 3

A student dilutes 25.0 cm^3 of $0.500 \text{ mol dm}^{-3}$ propanoic acid by adding water until the total volume is 100.0 cm^3

Calculate the pH of this diluted solution of propanoic acid.

Give your answer to 2 decimal places.

[4 marks]

$$[\text{C}_2\text{H}_5\text{COOH}] = 0.500 \times 25.0 \div 100.0 = 0.125 \text{ mol dm}^{-3}$$

$$K_a = [\text{H}^+]^2 \div [\text{C}_2\text{H}_5\text{COOH}]$$

$$[\text{H}^+] = \sqrt{K_a \times [\text{HA}]} = \sqrt{1.35 \times 10^{-5} \times 0.125}$$

$$[\text{H}^+] = 1.30 \times 10^{-3} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log_{10}(1.30 \times 10^{-3}) = 2.89$$

pH 2.89



0 4 . 4

A buffer solution with a pH of 4.50 is made by dissolving x g of sodium propanoate ($\text{C}_2\text{H}_5\text{COONa}$) in a solution of propanoic acid. The final volume of buffer solution is 500 cm^3 and the final concentration of the propanoic acid is $0.250 \text{ mol dm}^{-3}$

Calculate x in g

For propanoic acid, $K_a = 1.35 \times 10^{-5} \text{ mol dm}^{-3}$

[6 marks]

$$[\text{H}^+] = 10^{-4.50} = 3.16 \times 10^{-5} \text{ mol dm}^{-3}$$

$$K_a = [\text{H}^+][\text{C}_2\text{H}_5\text{COO}^-] \div [\text{C}_2\text{H}_5\text{COOH}]$$

$$[\text{C}_2\text{H}_5\text{COO}^-] = K_a \times [\text{C}_2\text{H}_5\text{COOH}] \div [\text{H}^+]$$

$$= (1.35 \times 10^{-5} \times 0.250) \div (3.16 \times 10^{-5})$$

$$[\text{C}_2\text{H}_5\text{COO}^-] = 0.1068 \text{ mol dm}^{-3}$$

$$M_r(\text{C}_2\text{H}_5\text{COONa}) = 3(12.0) + 5(1.0) + 2(16.0) + 23.0 = 96.0$$

$$n(\text{C}_2\text{H}_5\text{COONa}) = 0.1068 \times 500 \div 1000 = 0.0534 \text{ mol}$$

$$x = \text{mass} = 0.0534 \times 96.0 = 5.13 \text{ g}$$

x 5.13 g

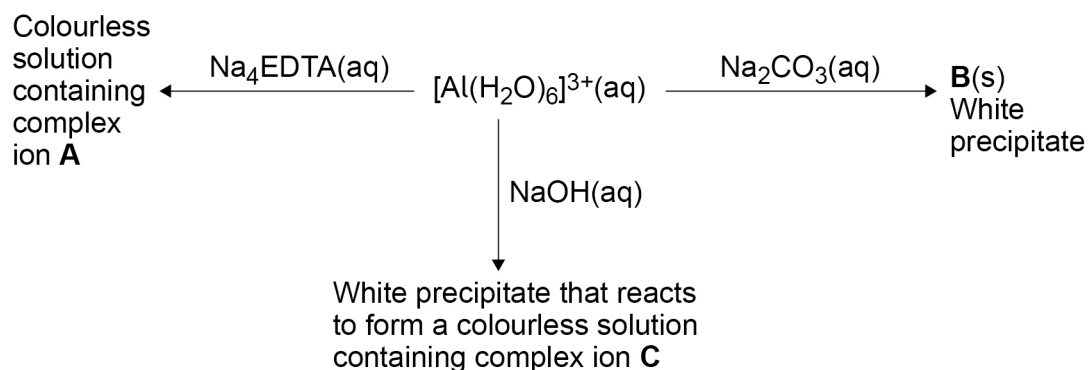
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0 5

Some reactions of the $[\text{Al}(\text{H}_2\text{O})_6]^{3+}(\text{aq})$ ion are shown.



0 5 . 1

Give the formula of the white precipitate **B**.

State **one** other observation when $\text{Na}_2\text{CO}_3(\text{aq})$ is added to a solution containing $[\text{Al}(\text{H}_2\text{O})_6]^{3+}(\text{aq})$ ions.

Give an equation for this reaction.

[3 marks]

Formula of **B** $\text{Al}(\text{H}_2\text{O})_3(\text{OH})_3$

Observation Bubbles of gas (effervescence)

Equation



0 5 . 2

Give the formula of the complex ion **C**.

State **one** condition needed for the formation of **C** from $[\text{Al}(\text{H}_2\text{O})_6]^{3+}(\text{aq})$ and $\text{NaOH}(\text{aq})$.

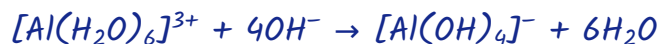
Give an equation for this reaction.

[3 marks]

Formula of **C** $[\text{Al}(\text{OH})_4]^-$

Condition Excess $\text{NaOH}(\text{aq})$

Equation



0 5 . 3 Deduce the formula of the complex ion **A**.

[1 mark]



0 5 . 4 Explain, with the use of an equation, why a solution containing $[Al(H_2O)_6]^{3+}$ has a pH < 7

[3 marks]

Equation



Explanation *The Al^{3+} ion is small and has a high charge, so it has a high charge density.*

It strongly polarises the water ligands, attracting electron density from the O-H bonds, weakening them so H^+ ions are released, so $[H^+]$ increases.

10

Turn over for the next question

Turn over ►



0 6

Methanol can be manufactured in a reversible reaction as shown.

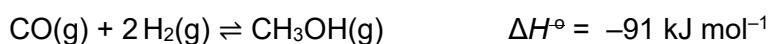
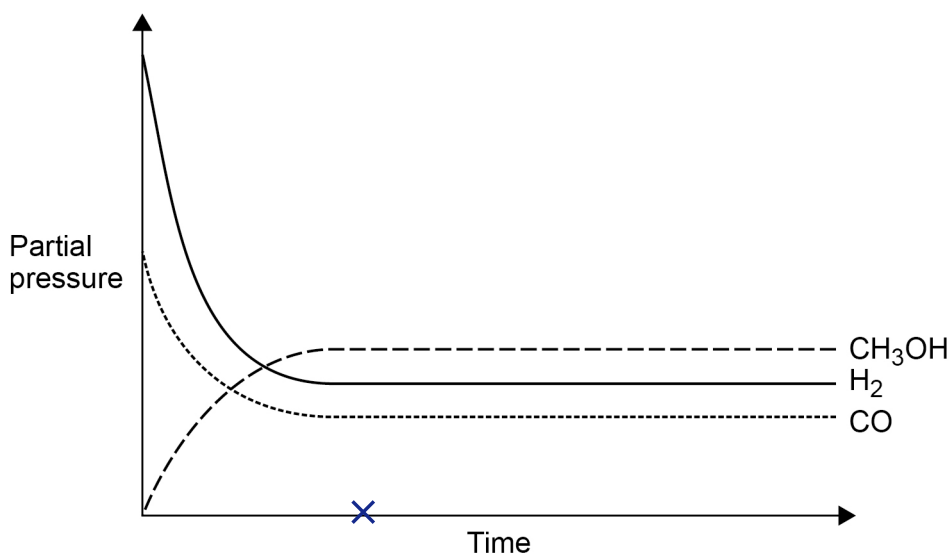


Figure 3 shows how the partial pressures change with time at a constant temperature.

Figure 3



0 6

1

Draw a cross (x) on the appropriate axis of **Figure 3** when the mixture reaches equilibrium.

[1 mark]

0 6

2

A 0.230 mol sample of carbon monoxide is mixed with hydrogen in a 1:2 mol ratio and allowed to reach equilibrium in a sealed flask at temperature T . At equilibrium the mixture contains 0.120 mol of carbon monoxide. The total pressure of this mixture is 1.04×10^4 kPa

Calculate the partial pressure, in kPa, of hydrogen in the equilibrium mixture.

[4 marks]

$$\text{CO reacted} = 0.230 - 0.120 = 0.110 \text{ mol, so } 0.220 \text{ mol } \text{H}_2 \text{ reacted}$$

$$\text{Equilibrium: } \text{CO} = 0.120, \text{H}_2 = 0.460 - 0.220 = 0.240, \text{CH}_3\text{OH} = 0.110 \text{ mol}$$

$$\text{Total moles} = 0.120 + 0.240 + 0.110 = 0.470 \text{ mol}$$

$$\text{Mole fraction of } \text{H}_2 = 0.240 \div 0.470 = 0.511$$

$$p(\text{H}_2) = 0.511 \times 1.04 \times 10^4 = 5310 \text{ kPa}$$

Partial pressure of hydrogen 5310 kPa



0 6 . 3 Give an expression for the equilibrium constant (K_p) for this reaction.

State the units.

[2 marks]

$$K_p = \frac{p(\text{CH}_3\text{OH})}{p(\text{H}_2)^2 \times p(\text{CO})}$$

Units kPa^{-2}

0 6 . 4 Some more carbon monoxide is added to the mixture in Question **06.2**. The new mixture is allowed to reach equilibrium at temperature T .

State the effect, if any, on the partial pressure of methanol and on the value of K_p

[2 marks]

Effect on partial pressure of methanol Increases

Effect on value of K_p No effect (stays the same)

0 6 . 5 State the effect, if any, of the addition of a catalyst on the value of K_p for this equilibrium.
Explain your answer.

[2 marks]

Effect on value of K_p No effect

Explanation A catalyst increases the rates of the forward and backward reactions equally, so the position of equilibrium does not change. Only temperature affects K_p .

11

Turn over for the next question

Turn over ►



0 7

The melting point of XeF_4 is higher than the melting point of PF_3

Explain why the melting points of these two compounds are different.

In your answer you should give the shape of each molecule, explain why each molecule has that shape and how the shape influences the forces that affect the melting point.

[6 marks]

In XeF_4 , xenon has 8 outer electrons. Four are used to form four bonds to fluorine atoms, leaving two lone pairs, so there are 4 bonding pairs and 2 lone pairs around Xe.

These six electron pairs repel each other as far apart as possible. Lone pairs repel more than bonding pairs, so the two lone pairs sit opposite each other (at 180°), above and below the plane. So XeF_4 is square planar with F-Xe-F bond angles of 90° .

In PF_3 , phosphorus has 5 outer electrons. Three are used to form three bonds to fluorine atoms, leaving one lone pair, so there are 3 bonding pairs and 1 lone pair around P.

The four electron pairs repel as far apart as possible (tetrahedral arrangement), but the lone pair repels more than the bonding pairs, so PF_3 is trigonal pyramidal with bond angles of about 107° .

F is much more electronegative than Xe and P, so all the bonds are polar. XeF_4 is symmetrical, so the bond dipoles cancel and the molecule is non-polar: it only has van der Waals forces between molecules.

PF_3 is pyramidal (not symmetrical), so the dipoles do not cancel and the molecule is polar: it has permanent dipole-dipole forces as well as van der Waals forces.

However, XeF_4 has a much larger M_r (207.3 compared with 88.0) and many more electrons, so it is a larger molecule and its van der Waals forces are much stronger than all the intermolecular forces in PF_3 .



So the intermolecular forces between XeF_4 molecules are stronger, more energy is needed to overcome them, and XeF_4 has the higher melting point.



0 8

A student does an experiment to determine the percentage by mass of sodium chlorate(I), NaClO, in a sample of bleach solution.

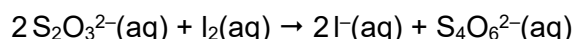
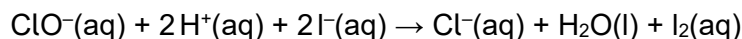
Method:

- Dilute a 10.0 cm³ sample of bleach solution to 100 cm³ with distilled water.
- Transfer 25.0 cm³ of the diluted bleach solution to a conical flask and acidify using sulfuric acid.
- Add excess potassium iodide to the conical flask to form a brown solution containing I₂(aq).
- Add 0.100 mol dm⁻³ sodium thiosulfate solution (Na₂S₂O₃) to the conical flask from a burette until the brown solution containing I₂(aq) becomes a colourless solution containing I⁻(aq).

The student uses 33.50 cm³ of sodium thiosulfate solution.

The density of the original bleach solution is 1.20 g cm⁻³

The equations for the reactions in this experiment are



0 8 . 1

Use all the information given to calculate the percentage by mass of NaClO in the original bleach solution.

Give your answer to 3 significant figures.

[7 marks]

$$n(\text{S}_2\text{O}_3^{2-}) = 33.50 \times 0.100 \div 1000 = 3.35 \times 10^{-3} \text{ mol}$$

$$n(\text{I}_2) = 3.35 \times 10^{-3} \div 2 = 1.675 \times 10^{-3} \text{ mol}$$

$$n(\text{ClO}^-) \text{ in } 25.0 \text{ cm}^3 = n(\text{I}_2) = 1.675 \times 10^{-3} \text{ mol}$$

$$n(\text{ClO}^-) \text{ in } 100 \text{ cm}^3 = 1.675 \times 10^{-3} \times 4 = 6.70 \times 10^{-3} \text{ mol}$$

$$= n(\text{NaClO}) \text{ in the original } 10.0 \text{ cm}^3 \text{ of bleach}$$

$$M_r(\text{NaClO}) = 74.5, \text{ mass} = 6.70 \times 10^{-3} \times 74.5 = 0.499 \text{ g}$$

$$\text{Mass of bleach} = 10.0 \times 1.20 = 12.0 \text{ g}$$

$$\% \text{ by mass} = (0.499 \div 12.0) \times 100 = 4.16\%$$

Percentage by mass 4.16%



0	8	.	2
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The total uncertainty from two readings and an end point error in using a burette is $\pm 0.15 \text{ cm}^3$

What is the total percentage uncertainty in using the burette in this experiment?

[1 mark]

Tick (✓) **one** box.

0.45%

☒

0.90%

☐

1.34%

☐

8

Turn over for the next question

Turn over ►



0 9

This question is about sodium halides.

0 9 . 1

State what is observed when silver nitrate solution is added to sodium fluoride solution.

[1 mark]

Colourless solution: no visible change (no precipitate).

0 9 . 2

State **one** observation when solid sodium chloride reacts with concentrated sulfuric acid.

Give an equation for the reaction.

State the role of the chloride ions in the reaction.

[3 marks]

Observation *Misty (steamy) white fumes*

Equation

Role *Base (proton acceptor)*

0 9 . 3

Give an equation for the redox reaction between solid sodium bromide and concentrated sulfuric acid.

Explain, using oxidation states, why this is a redox reaction.

[3 marks]

Equation

Explanation *Br changes oxidation state from -1 (in Br⁻) to 0 (in Br₂),
so bromide is oxidised.**S changes oxidation state from +6 (in H₂SO₄) to +4 (in SO₂), so
sulfur is reduced.*

0 9 . 4

State what is observed when aqueous chlorine is added to sodium bromide solution.

Give an ionic equation for the reaction.

[2 marks]

Observation *Solution turns yellow / orange*

Ionic equation



9



1 0

Methanol is formed when carbon dioxide and hydrogen react.

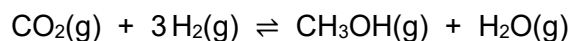


Table 5 contains enthalpy of formation and entropy data for these substances.

Table 5

	CO₂(g)	H₂(g)	CH₃OH(g)	H₂O(g)
$\Delta_f H / \text{kJ mol}^{-1}$	-394	0	-201	-242
$S / \text{J K}^{-1} \text{mol}^{-1}$	214	131	238	189

1 0

1

Use the equation and the data in **Table 5** to calculate the Gibbs free-energy change (ΔG), in kJ mol^{-1} , for this reaction at 890 K

[6 marks]

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

$$\Delta H = (-201 + -242) - (-394 + 0) = -49 \text{ kJ mol}^{-1}$$

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

$$\Delta S = (238 + 189) - (214 + 3 \times 131) = 427 - 607$$

$$\Delta S = -180 \text{ J K}^{-1} \text{mol}^{-1} = -0.180 \text{ kJ K}^{-1} \text{mol}^{-1}$$

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

$$\Delta G = -49 - (890 \times -0.180) = -49 + 160.2$$

$$\Delta G = +111 \text{ kJ mol}^{-1}$$

$$\Delta G \quad \underline{\quad +111 \quad} \quad \text{kJ mol}^{-1}$$

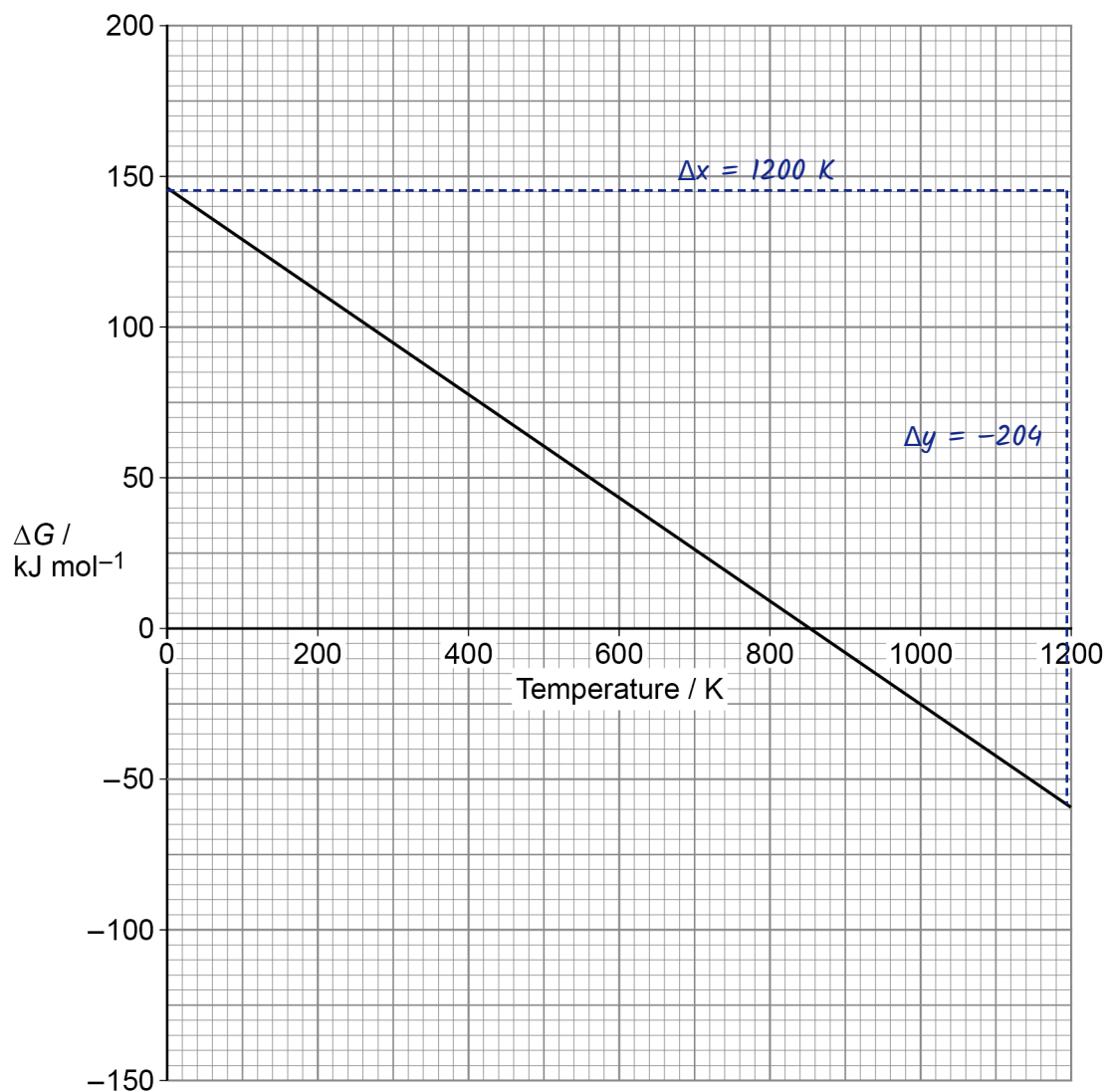
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Figure 4 shows how the Gibbs free-energy change varies with temperature in a different gas phase reaction.

The straight line graph for this gas phase reaction has been extrapolated to zero Kelvin.

Figure 4



1 0 . 2

Use the values of the intercept and gradient from the graph in **Figure 4** to calculate the enthalpy change (ΔH), in kJ mol^{-1} , and the entropy change (ΔS), in $\text{J K}^{-1} \text{mol}^{-1}$, for this reaction.

[4 marks]

$$\Delta G = \Delta H - T\Delta S, \text{ so the intercept (at } T = 0 \text{ K)} = \Delta H$$

$$\Delta H = +145 \text{ kJ mol}^{-1}$$

$$\text{The gradient of the line} = -\Delta S$$

$$\text{gradient} = (-59 - 145) \div (1200 - 0) = -0.170 \text{ kJ K}^{-1} \text{mol}^{-1}$$

$$\Delta S = -\text{gradient} = +0.170 \text{ kJ K}^{-1} \text{mol}^{-1}$$

$$\Delta S = 0.170 \times 1000 = +170 \text{ J K}^{-1} \text{mol}^{-1}$$

$$\Delta H \quad \underline{\quad +145 \quad} \quad \text{kJ mol}^{-1}$$

$$\Delta S \quad \underline{\quad +170 \quad} \quad \text{J K}^{-1} \text{mol}^{-1}$$

1 0 . 3

State what **Figure 4** shows about the feasibility of the reaction.

[1 mark]

ΔG is negative above about 850 K, so the reaction is only
(thermodynamically) feasible above 850 K; below 850 K it is not
feasible.

Turn over ►



1 1

This question is about a glucose–oxygen fuel cell.

When the cell operates, the glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) molecules react with water at the negative electrode to form carbon dioxide and hydrogen ions.

Oxygen gas reacts with hydrogen ions to form water at the positive electrode.

1 1 . 1

Deduce the half-equation for the reaction at the negative electrode.

[1 mark]



1 1 . 2

Deduce the half-equation for the reaction at the positive electrode.

[1 mark]



1 1 . 3

Give the equation for the overall reaction that occurs in the Glucose–oxygen fuel cell.

[1 mark]



1 1 . 4

The negative electrode is made of carbon and the positive electrode is made of platinum.

Give the conventional representation for the glucose–oxygen fuel cell.

[2 marks]



1 1 . 5

State what must be done to maintain the EMF of this fuel cell when in use.

[1 mark]

Constantly add glucose (and oxygen) to the cell.

6

END OF QUESTIONS



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3 2



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