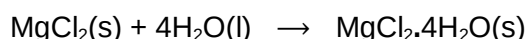


Name of the Student: _____

Max. Marks : 26 Marks

Time : 26 Minutes

Q1.

Anhydrous magnesium chloride, MgCl_2 , can absorb water to form the hydrated salt $\text{MgCl}_2 \cdot 4\text{H}_2\text{O}$ 

- (a) Suggest **one** reason why the enthalpy change for this reaction cannot be determined directly by calorimetry.

(1)

- (b) Some enthalpies of solution are shown in Table 1.

Table 1

Salt	Enthalpy of solution / kJ mol^{-1}
$\text{MgCl}_2(\text{s})$	-155
$\text{MgCl}_2 \cdot 4\text{H}_2\text{O}(\text{s})$	-39

Calculate the enthalpy change for the absorption of water by $\text{MgCl}_2(\text{s})$ to form $\text{MgCl}_2 \cdot 4\text{H}_2\text{O}(\text{s})$.Enthalpy change _____ kJ mol^{-1}

(2)

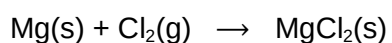
- (c) Describe how you would carry out an experiment to determine the enthalpy of solution of anhydrous magnesium chloride.

You should use about 0.8 g of anhydrous magnesium chloride.

Explain how your results could be used to calculate the enthalpy of solution.

(6)

- (d) Anhydrous magnesium chloride can be formed by direct reaction between its elements.



The free-energy change, ΔG , for this reaction varies with temperature as shown in **Table 2**.

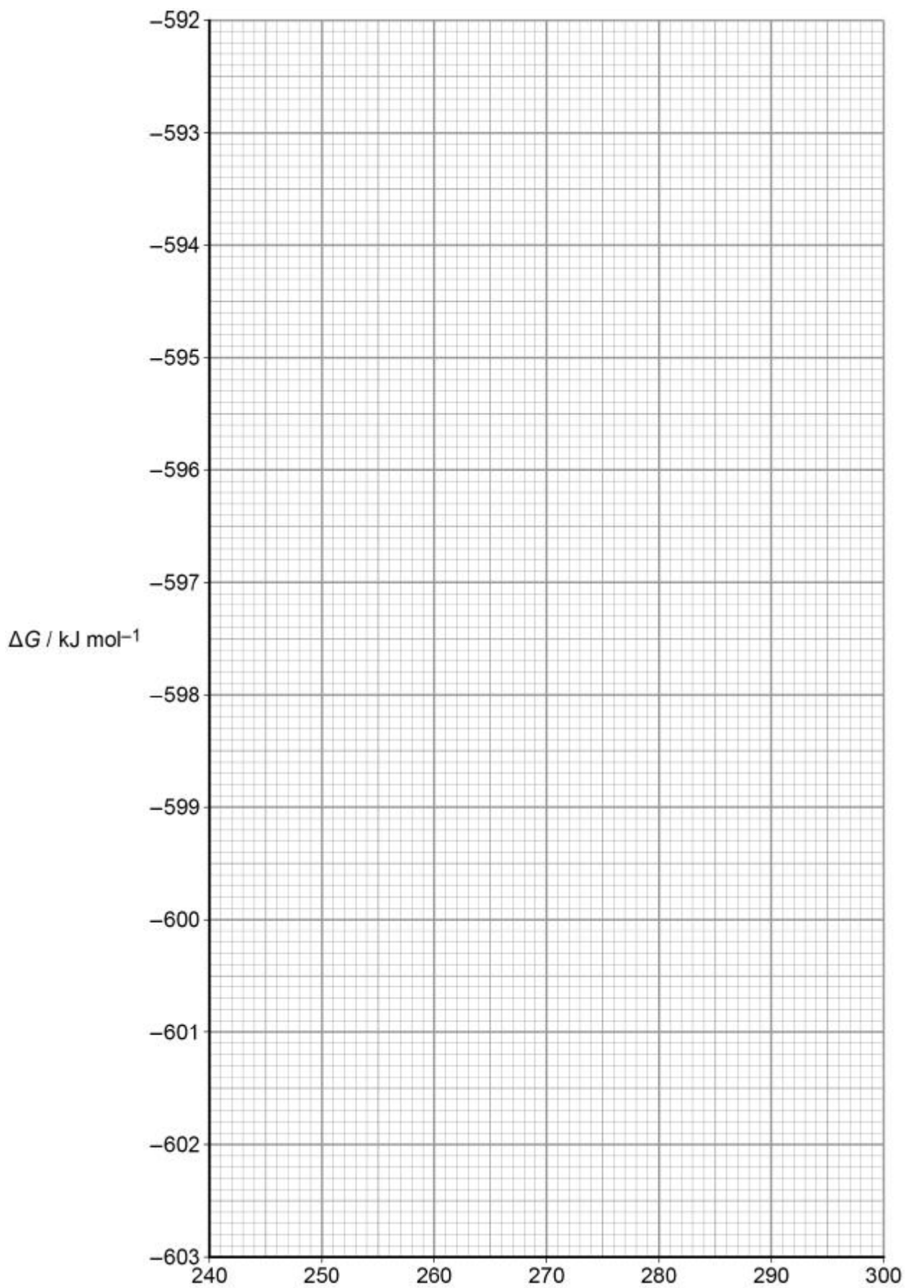
Table 2

T / K	$\Delta G / \text{kJ mol}^{-1}$
298	-592.5
288	-594.2
273	-596.7
260	-598.8
240	-602.2

Use these data to plot a graph of free-energy change against temperature on the grid below.

Calculate the gradient of the line on your graph and hence calculate the entropy change, ΔS , in $\text{J K}^{-1} \text{mol}^{-1}$, for the formation of anhydrous magnesium chloride from its elements.

Show your working.



ΔS _____ $\text{J K}^{-1} \text{mol}^{-1}$

(5)

(Total 14 marks)

Q2.

A reaction is exothermic and has a negative entropy change.

Which statement is correct?

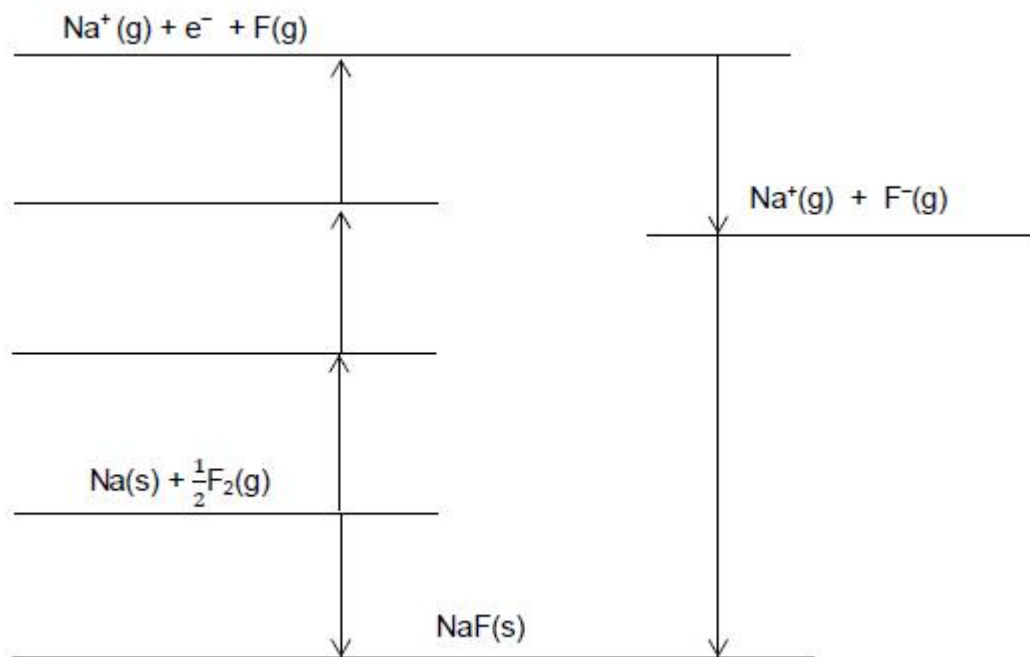
- | | | |
|----------|--|--------------------------|
| A | The reaction is always feasible | ☐ |
| B | The reaction is feasible above a certain temperature | ☐ |
| C | The reaction is feasible below a certain temperature | ☐ |
| D | The reaction is never feasible | ☐ |

(Total 1 mark)

Q3.

This question is about sodium fluoride.

- (a) Complete the Born–Haber cycle for sodium fluoride by adding the missing species on the lines.



(2)

- (b) Use the data in the table and your completed Born–Haber cycle from part (a) to calculate the enthalpy of lattice formation of sodium fluoride.

	$\Delta H^\circ / \text{kJ mol}^{-1}$
$\text{Na(s)} \rightarrow \text{Na(g)}$	+109
$\text{Na(g)} \rightarrow \text{Na}^+(\text{g}) + \text{e}^-$	+494
$\text{F}_2(\text{g}) \rightarrow 2\text{F(g)}$	+158
$\text{F(g)} + \text{e}^- \rightarrow \text{F}^-(\text{g})$	-348
$\text{Na(s)} + \frac{1}{2}\text{F}_2(\text{g}) \rightarrow \text{NaF(s)}$	-569

Enthalpy of lattice formation _____ kJ mol^{-1}

(2)

- (c) Suggest how the enthalpy of lattice formation of NaCl compares with that of NaF

Justify your answer.

How enthalpies of formation compare _____

Justification _____

(3)

- (d) Calculate the volume, in cm^3 , of fluorine gas at 298 K and 100 kPa required to produce 1.00 g of sodium fluoride by reaction with an excess of sodium.

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

Volume _____ cm^3

(4)

(Total 11 marks)