Q1.

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

(a) State the meaning of the term enthalpy change.

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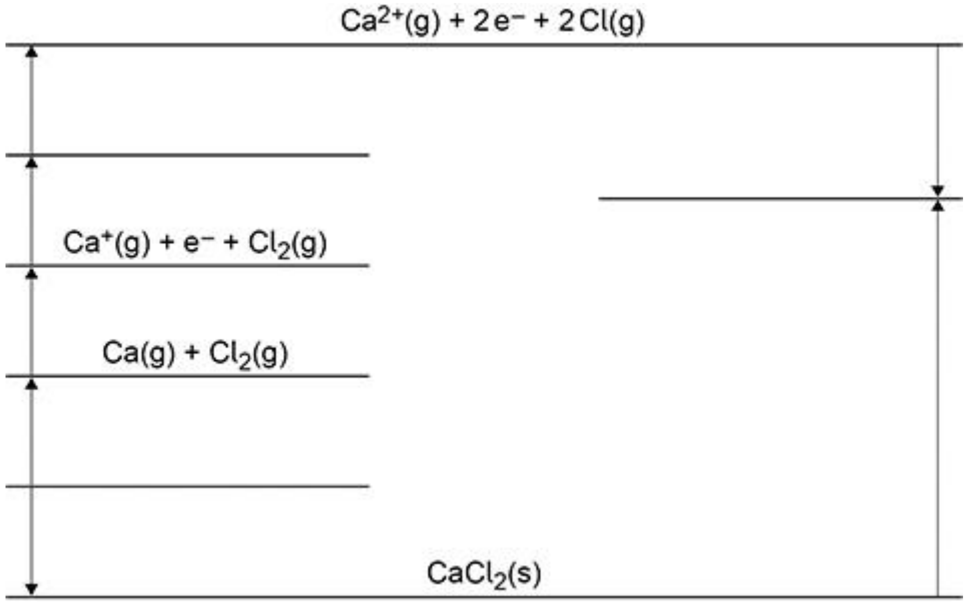
(1)

The figure below shows an incomplete Born–Haber cycle for the formation of calcium chloride.



(b) Complete the figure above by writing the formulas, including state symbols, of the appropriate species on each of the three blank lines.

(3)

(c) **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 the figure in part (a) and the data in **Table 1** to calculate a value for the enthalpy of lattice dissociation of calcium chloride.

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

(2)

(d) 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.

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(1)

(e) **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 Mg^{2+}(g)	–1920
Enthalpy of hydration of Cl^-(g)	–364

Use your answer to part (d) and the data in **Table 2** to calculate a value for the enthalpy of solution of magnesium chloride.

Enthalpy of solution _______________ kJ mol^{-1}

(2)

(f) The enthalpy of hydration of Ca^{2+}(g) is –1650 kJ mol^{-1}

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

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(2)

(Total 11 marks)

Q2.

This question is about thermodynamics.
Consider the reaction shown.

$$2\ Al_2O_3(s) + 3\ C(s) \rightarrow 4\ Al(s) + 3\ CO_2(g)$$

The table below shows some thermodynamic data.

Substance	**Al_2O_3(s)**	**Al(s)**	**C(s)**	**CO_2(g)**
$\Delta_f H^\ominus$/ kJ mol^{-1}	–1669	0	0	–394
$S^\ominus$/ J K^{-1} mol^{-1}	51	28	6	214

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

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(1)

(b) State the temperature at which the standard entropy of aluminium is 0 J K^{-1} mol^{-1}

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(1)

(c) Use the equation and the data in the table above to calculate the minimum temperature, in K, at which this reaction becomes feasible.

Minimum temperature _______________ K

(7)

(Total 9 marks)

Q3.

This question is about enthalpy changes.

(a) The figure below shows a Born–Haber cycle for the formation of strontium chloride, $SrCl_2$

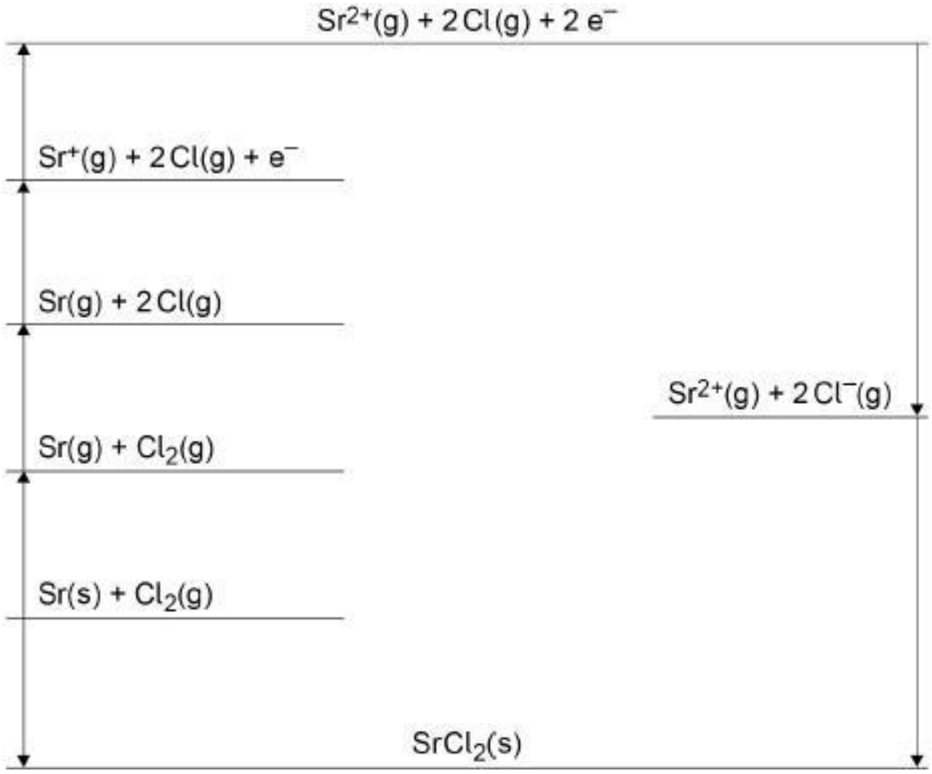

Table 1 shows some thermodynamic data.

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.

Electron affinity _______________ kJ mol^{-1}

(3)

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

Substance	**Enthalpy of lattice formation / kJ mol^{-1}**
$MgCl_2$	–2018
MgO	–2493
$BaCl_2$	–3889

(1)

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

(c) State why there is a difference between the theoretical and experimental values.

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(1)

(d) **Table 3** shows enthalpy of hydration values for ions of some Group 1 elements.

Table 3

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

Explain why the enthalpy of hydration becomes less exothermic from Li^+ to K^+

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(2)

(e) 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.

Enthalpy of hydration of bromide ions _______________ kJ mol^{-1}

(3)

(Total 10 marks)

Q4.

Methanol is formed when carbon dioxide and hydrogen react.

$$CO_2(g) + 3\ H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

The table contains enthalpy of formation and entropy data for these substances.

	CO_2(g)	**H_2(g)**	**CH_3OH(g)**	**H_2O(g)**
$\Delta_f H$ / kJ mol^{-1}	–394	0	–201	–242
S / J K^{-1} mol^{-1}	214	131	238	189

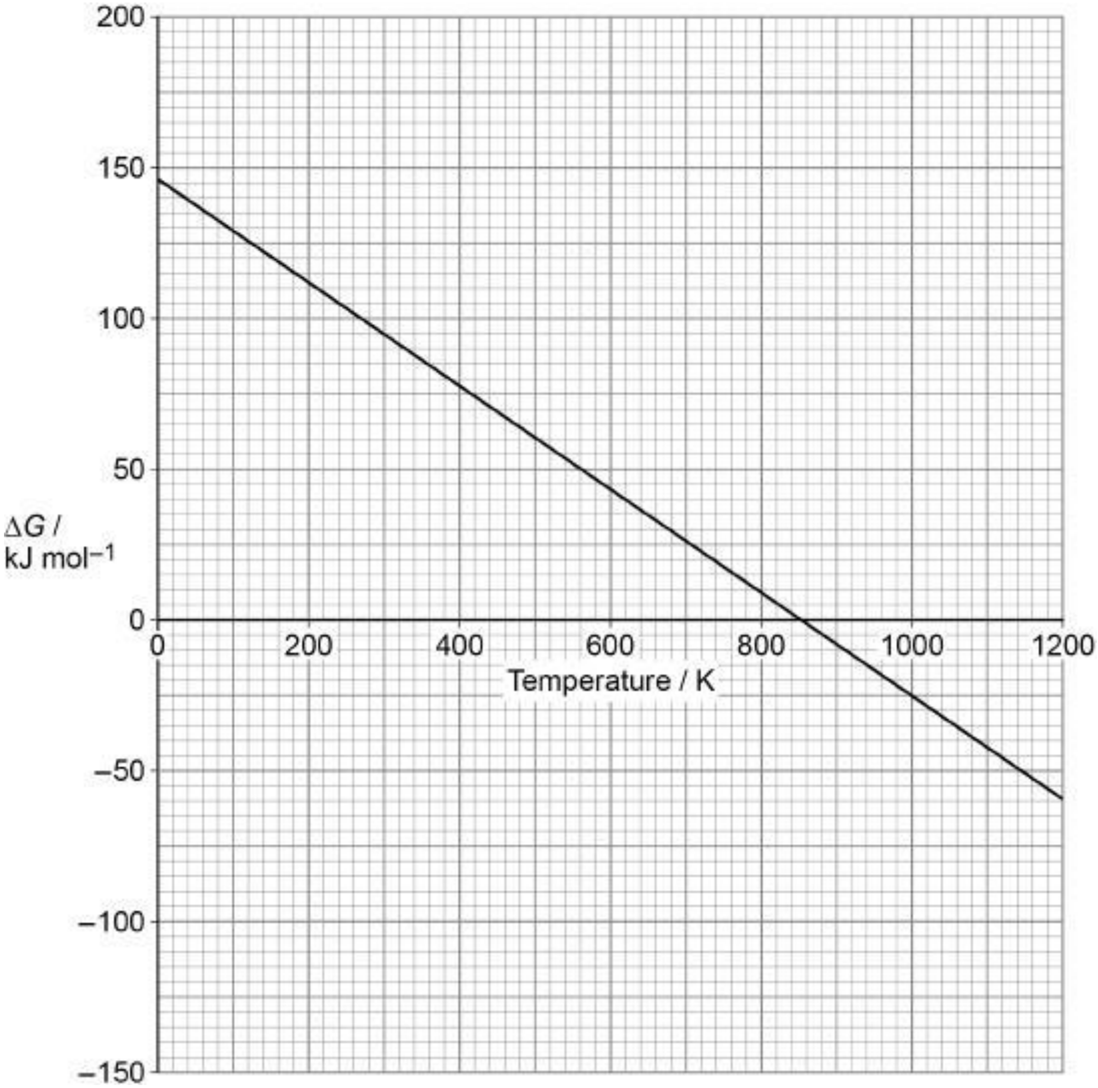
(a) Use the equation and the data in the table above to calculate the Gibbs free-energy change (ΔG), in kJ mol^{-1}, for this reaction at 890 K

ΔG _______________ kJ mol^{-1}

(6)

The graph below 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.



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

ΔH _______________ kJ mol^{-1}

ΔS _______________ J K^{-1} mol^{-1}

(4)

(c) State what the graph above shows about the feasibility of the reaction.

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(1)

(Total 11 marks)

Q5.

The diagram shows an incomplete Born−Haber cycle for the formation of caesium iodide. The diagram is not to scale.

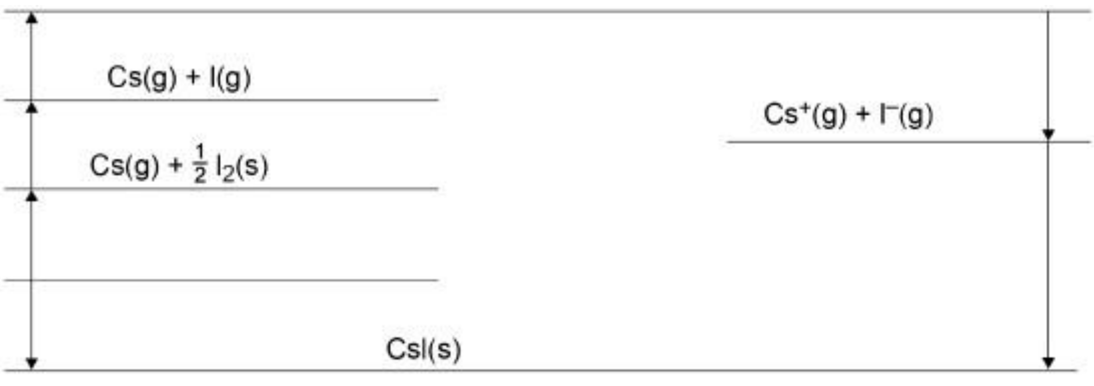

Table 1 gives values of some standard enthalpy changes.

Table 1

Name of enthalpy change	**$\Delta H^{\ominus}$ / kJ mol^{-1}**
Enthalpy of atomisation of caesium	+79
First ionisation energy of caesium	+376
Electron affinity of iodine	−314
Enthalpy of lattice formation of caesium iodide	−585
Enthalpy of formation of caesium iodide	−337

(a) Complete the diagram above by writing the formulas, including state symbols, of the appropriate species on each of the two blank lines.

(2)

(b) Use the diagram above and the data in **Table 1** to calculate the standard enthalpy of atomisation of iodine.

Standard enthalpy of atomisation of iodine ___________________________ kJ mol^{-1}

(2)

(c) The enthalpy of lattice formation for caesium iodide in **Table 1** is a value obtained by experiment.
The value obtained by calculation using the perfect ionic model is −582 kJ mol^{-1}

Deduce what these values indicate about the bonding in caesium iodide.

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(1)

(d) Use data from **Table 2** to show that this reaction is **not** feasible at 298 K

$$CsI(s) \rightarrow Cs(s) + \frac{1}{2}I_2(s) \qquad \Delta H^\ominus = +337 \text{ kJ mol}^{-1}$$

Table 2

	CsI(s)	**Cs(s)**	**I_2(s)**
$S^\ominus$ / J K^{-1} mol^{-1}	130	82.8	117

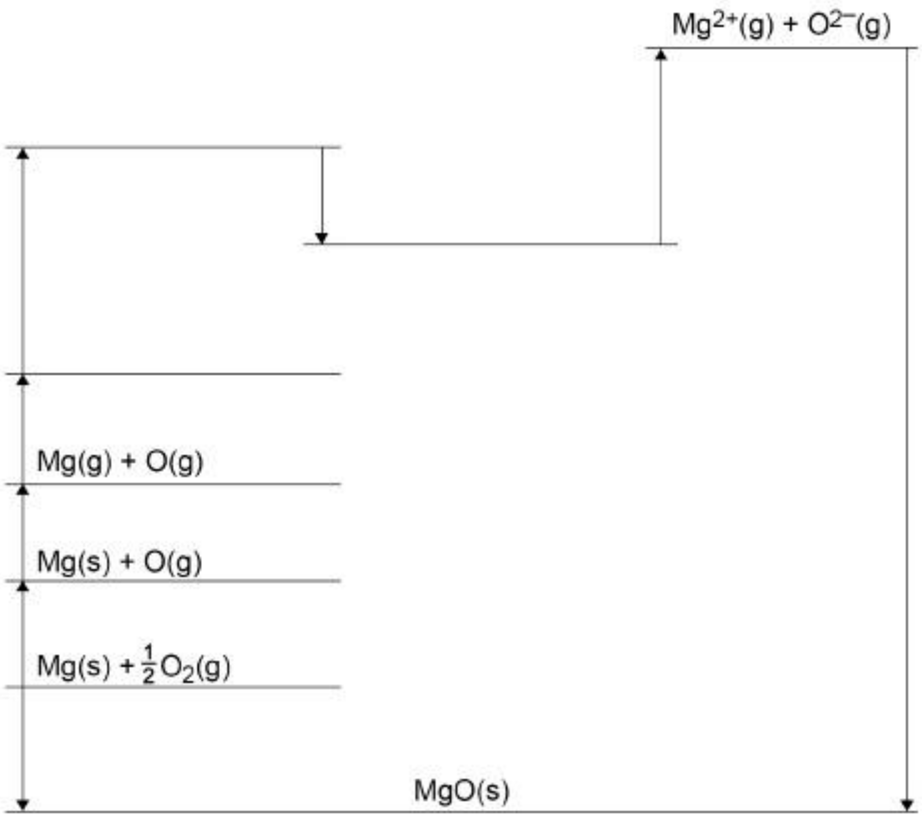
(4)
(Total 9 marks)

Q6.

This question is about lattice enthalpies.

(a) The diagram shows a Born–Haber cycle for the formation of magnesium oxide.

Complete the diagram by writing the missing symbols on the appropriate energy levels.



(3)

(b) The table contains some thermodynamic data.

	Enthalpy change / kJ mol^{-1}
Enthalpy of formation for magnesium oxide	–602
Enthalpy of atomisation for magnesium	+150
First ionisation energy for magnesium	+736
Second ionisation energy for magnesium	+1450
Bond dissociation enthalpy for oxygen	+496
First electron affinity for oxygen	–142
Second electron affinity for oxygen	+844

Calculate a value for the enthalpy of lattice formation for magnesium oxide.

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

(3)

(Total 6 marks)

Q7.

Which equation represents the process that occurs when the standard enthalpy of atomisation of iodine is measured?

A $\frac{1}{2} I_2(s) \rightarrow I(g)$ ☐

B $I_2(s) \rightarrow 2I(g)$ ☐

C $\frac{1}{2} I_2(g) \rightarrow I(g)$ ☐

D $I_2(g) \rightarrow 2I(g)$ ☐

(Total 1 mark)

Q8.

Lattice enthalpy values can be obtained from Born–Haber cycles and by calculations based on a perfect ionic model.

Which compound shows the greatest percentage difference between these two values?

A CsF ☐

B CsI ☐

C LiF ☐

D LiI ☐

(Total 1 mark)

Q9.

This question is about silver iodide.

(a) Define the term enthalpy of lattice formation.

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(2)

(b) Some enthalpy change data are shown in the table.

	Enthalpy change / kJ mol^{-1}
$AgI(s) \longrightarrow Ag^{+}(aq) + I^{-}(aq)$	+112
$Ag^{+}(g) \longrightarrow Ag^{+}(aq)$	−464
$I^{-}(g) \longrightarrow I^{-}(aq)$	−293

Use the data in the table to calculate the enthalpy of lattice formation of silver iodide.

Enthalpy of lattice formation = ____________________ kJ mol^{-1}

(2)

(c) A calculation of the enthalpy of lattice formation of silver iodide based on a perfect ionic model gives a smaller numerical value than the value calculated in part (b)

Explain this difference.

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(2)

(d) Identify a reagent that could be used to indicate the presence of iodide ions in an aqueous solution and describe the observation made.

Reagent ____

Observation ____

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(2)
(Total 8 marks)

Q10.

Titanium(IV) chloride can be made from titanium(IV) oxide as shown in the equation.

$TiO_2(s) + 2C(s) + 2Cl_2(g) \rightarrow 2CO(g) + TiCl_4(l)$ $\Delta H^\ominus = -60.0\ \text{kJ mol}^{-1}$

Some entropy data are shown in the table.

Substance	$TiO_2(s)$	C(s)	$Cl_2(g)$	CO(g)	$TiCl_4(l)$
$S^\ominus$ / J K^{-1} mol^{-1}	50.2	5.70	223	198	253

Use the equation and the data in the table to calculate the Gibbs free-energy change for this reaction at 989 °C

Give your answer to the appropriate number of significant figures.

Use your answer to explain whether this reaction is feasible.

Gibbs free-energy change ____________________ kJ mol^{-1}

Explanation

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(Total 6 marks)

Q11.

Anhydrous magnesium chloride, $MgCl_2$, can absorb water to form the hydrated salt $MgCl_2.4H_2O$

$$MgCl_2(s) + 4H_2O(l) \longrightarrow MgCl_2.4H_2O(s)$$

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

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(1)

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

Table 1

Salt	**Enthalpy of solution / kJ mol^{-1}**
$MgCl_2(s)$	−155
$MgCl_2.4H_2O(s)$	−39

Calculate the enthalpy change for the absorption of water by $MgCl_2(s)$ to form $MgCl_2.4H_2O(s)$.

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.

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(6)

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

$$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$$

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

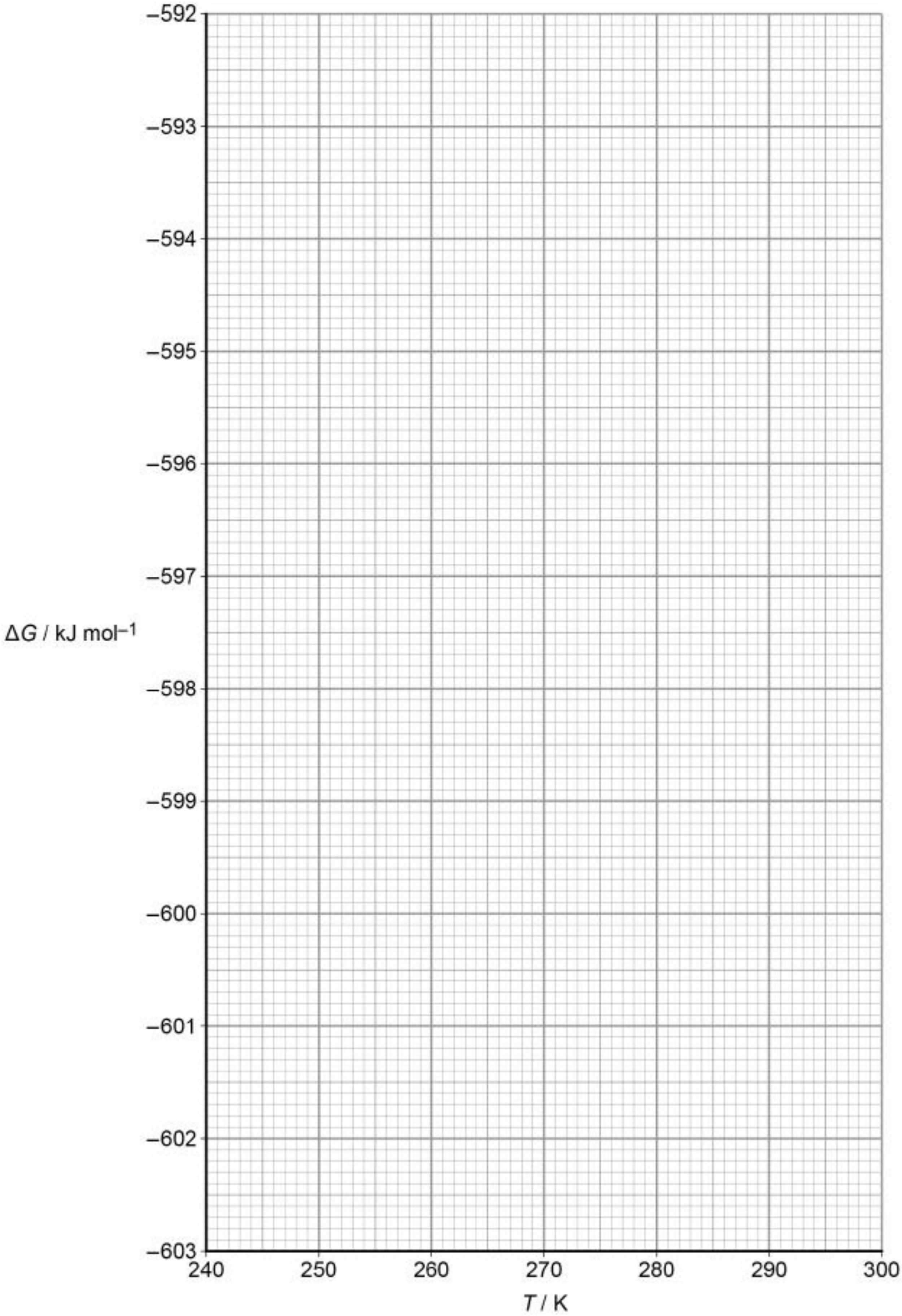
Table 2

T / K	**ΔG / 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 J K^{-1} mol^{-1}, for the formation of anhydrous magnesium chloride from its elements.

Show your working.


−592
−593
−594
−595
−596
−597
ΔG / kJ mol−1
−598
−599
−600
−601
−602
−603
240
250
260
270
280
290
300
T / K

ΔS ______________________________ $J\,K^{-1}\,mol^{-1}$

(5)

(Total 14 marks)

Q12.

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)