Q1.

This question is about enthalpy changes.

(a) Theoretical values for enthalpies of lattice dissociation can be calculated using a perfect ionic model.

State the meaning of the term perfect ionic model.

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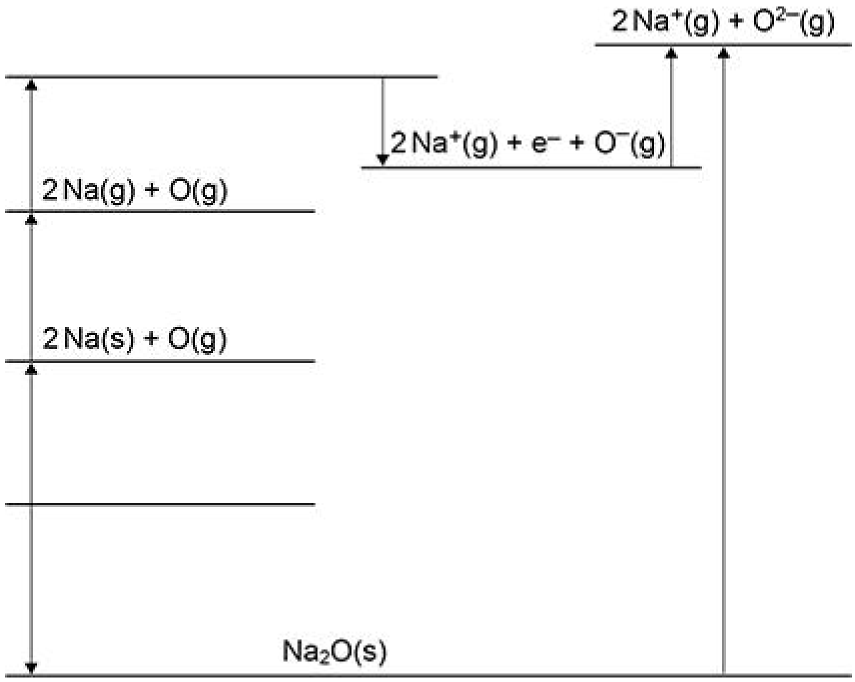
(1)

(b) Enthalpies of lattice dissociation can also be obtained from Born–Haber cycles.

The figure below shows an incomplete Born–Haber cycle for the formation of sodium oxide.



Complete above figure by writing formulas, including state symbols, of the appropriate species on each of the two blank lines.

(2)

(c) **Table 1** shows some enthalpy changes.

Table 1

Enthalpy change	**Δ*H* / kJ mol^{-1}**
Enthalpy of atomisation of oxygen	+248
Enthalpy of atomisation of sodium	+109
Enthalpy of formation of sodium oxide	–416
First ionisation energy of sodium	+494
First electron affinity of oxygen	–142
Second electron affinity of oxygen	+844

Use the data in **Table 1** to calculate the enthalpy of lattice dissociation of sodium oxide.

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

(2)

(d) Explain why the second electron affinity of oxygen has a positive value.

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(1)

(e) Explain why the enthalpy of lattice dissociation for sodium oxide is greater than the enthalpy of lattice dissociation for sodium chloride.

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(2)

(f) Sodium chloride dissolves in water.

Table 2 shows some more enthalpy changes.

Table 2

Enthalpy change	**Δ*H* / kJ mol^{-1}**
Enthalpy of hydration for Cl^- ions	–364
Enthalpy of hydration for Na^+ ions	–406
Enthalpy of lattice dissociation for NaCl	+771

Use the data in **Table 2** to calculate the enthalpy of solution for sodium chloride.

Enthalpy of solution ____________________ kJ mol^{-1}

(2)

(g) Give a reason why data books do **not** contain a value for the enthalpy of solution of sodium oxide.

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(1)

(h) Calculate the temperature, in °C, at which this reaction becomes feasible.

$$NaCl(s) \rightarrow Na(s) + \frac{1}{2} Cl_2(g) \qquad \Delta H = +411 \text{ kJ mol}^{-1}$$

$$\Delta S = +90.1 \text{ J K}^{-1} \text{ mol}^{-1}$$

Temperature ____________________ °C

(3)

(Total 14 marks)

Q2.

This question is about metal chlorides.

(a) **Table 1** shows some enthalpy change data.

Table 1

	Enthalpy change / kJ mol^{-1}
$Ca^{2+}(g) \rightarrow Ca^{2+}(aq)$	–1650
$Cl^{-}(g) \rightarrow Cl^{-}(aq)$	–364
$Ca^{2+}(g) + 2\ Cl^{-}(g) \rightarrow CaCl_2(s)$	–2237

Use the data in **Table 1** to calculate the molar enthalpy change when calcium chloride dissolves in water.

Molar enthalpy change ____________________ kJ mol^{-1}

(2)

(b) Use your answer to part (a) to deduce how the temperature changes when calcium chloride dissolves in water.

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(1)

(c) Explain why the enthalpy of hydration of fluoride ions is more negative than the enthalpy of hydration of chloride ions.

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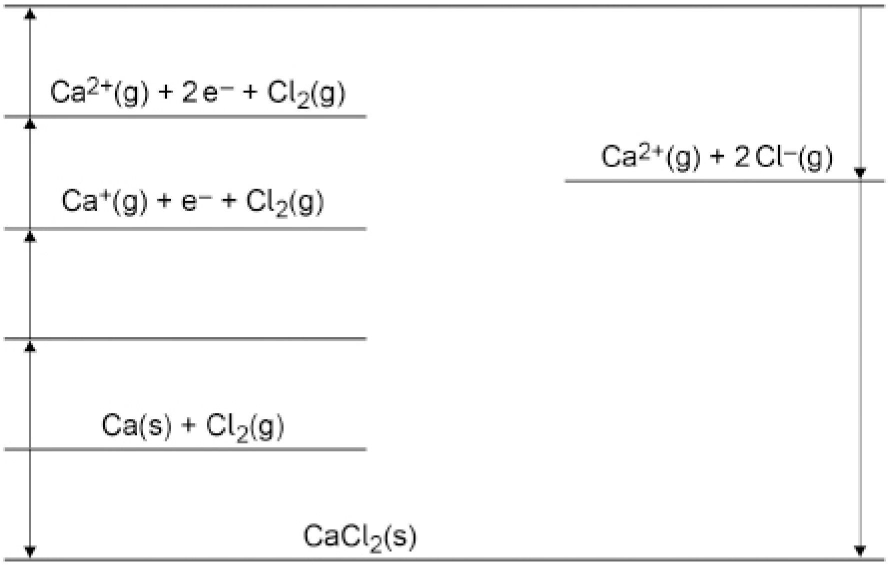
(2)

(d) The figure below shows an incomplete Born–Haber cycle for calcium chloride.



Complete the Born–Haber cycle by writing the formulas of the missing species on each of the two blank lines.

(2)

(e) **Table 2** shows some enthalpy change data.

Table 2

	Enthalpy change / kJ mol^{-1}
Enthalpy of atomisation of calcium	+193
First ionisation energy of calcium	+590
Enthalpy of atomisation of chlorine	+121
Electron affinity of chlorine	–364
Enthalpy of formation of calcium chloride	–795
Enthalpy of lattice formation of calcium chloride	–2237

Use the figure in (d) and data from **Table 2** to calculate the second ionisation energy of calcium.

Second ionisation energy ____________________ kJ mol^{-1}

(2)

(f) Explain why the second ionisation energy of calcium is greater than the first ionisation energy of calcium.

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(1)

(g) **Table 3** shows lattice enthalpies based on a perfect ionic model and lattice enthalpies from Born–Haber cycles for three metal chlorides.

Table 3

	Lattice enthalpy of dissociation / kJ mol^{-1}	
	Perfect ionic model	**Born–Haber cycle**
Calcium chloride	2223	2237
Potassium chloride	690	701
Silver chloride	770	905

Discuss the values in **Table 3**.

In your answer you should

- compare the three values based on a perfect ionic model
- compare the values based on a perfect ionic model to the values from a Born–Haber cycle for each compound.

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(6)
(Total 16 marks)

Q3.

A value for enthalpy of solution can be determined in two ways:

- from a cycle, using lattice enthalpy and enthalpies of hydration
- from the results of a calorimetry experiment.

(a) Define the term enthalpy of lattice dissociation.

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(2)

(b) The enthalpy of solution for ammonium nitrate is the enthalpy change for the reaction shown.

$NH_4NO_3(s) + aq \rightarrow NH_4^+(aq) + NO_3^-(aq) \quad \Delta H = +26 \text{ kJ mol}^{-1}$

Table 1

	$NH_4^+(g)$	$NO_3^-(g)$
Enthalpy of hydration $\Delta_{hyd}H$ / kJ mol⁻¹	−307	−314

Draw a suitably labelled cycle and use it, with data from **Table 1**, to calculate the enthalpy of lattice dissociation for ammonium nitrate.

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

(3)

(c) A student does an experiment to determine a value for the enthalpy of solution for ammonium nitrate.

The student uses this method.

- Measure 25.0 cm^3 of distilled water in a measuring cylinder.
- Pour the water into a beaker.
- Record the temperature of the water in the beaker.
- Add 4.00 g of solid NH_4NO_3 to the water in the beaker.
- Stir the solution and record the lowest temperature reached.

Table 2 shows the student's results.

Table 2

Initial temperature / °C	20.2
Lowest temperature / °C	12.2

Calculate the enthalpy of solution, in kJ mol^{-1}, for ammonium nitrate in this experiment.

Assume that the specific heat capacity of the solution, c = 4.18 J K^{-1} g^{-1}

Assume that the density of the solution = 1.00 g cm^{-3}

Enthalpy of solution _______________ kJ mol^{-1}

(3)

(d) The uncertainty in each of the temperature readings from the thermometer used in this experiment is ±0.1°C

Calculate the percentage uncertainty in the temperature change in this experiment.

Percentage uncertainty ____________________

(1)

(e) Suggest a change to the student's method, using the same apparatus, that would reduce the percentage uncertainty in the temperature change.

Give a reason for your answer.

Change ____

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Reason ____

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(2)

(f) Another student obtained a value of +15 kJ mol^{-1} using the same method.

Suggest the main reason for the difference between this experimental value for the enthalpy of solution and the correct value of +26 kJ mol^{-1}

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(1)

(g) **Table 3** shows some entropy data at 298 K

Table 3

	Entropy *S* / J K^{-1} mol^{-1}
$NH_4NO_3(s)$	151
$NH_4^+(aq)$	113
$NO_3^-(aq)$	146

Calculate a value for the Gibbs free-energy change (ΔG), at 298 K, for the reaction when ammonium nitrate dissolves in water.

$NH_4NO_3(s) + aq \rightarrow NH_4^+(aq) + NO_3^-(aq)$ $\Delta H = +26$ kJ mol^{-1}

Use data from **Table 3** and the value of ΔH from the equation.
Assume for the solvent, water, that the entropy change, $\Delta S = 0$

Explain what the calculated value of ΔG indicates about the feasibility of this reaction at 298 K

ΔG ________________ kJ mol^{-1}

Explanation ____

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(4)

(h) Ammonium nitrate decomposes as shown.

$NH_4NO_3(s) \rightarrow N_2(g) + \frac{1}{2} O_2(g) + 2 H_2O(g)$ $\Delta H = +123$ kJ mol^{-1}

The entropy change (ΔS) for this reaction is +144 J K^{-1} mol^{-1}

Calculate the temperature at which this reaction becomes feasible.

Temperature ________________ K

(2)

(Total 18 marks)