2 **(a)** Predict and explain the variation in enthalpy change of hydration for the ions Na^+, Mg^{2+} and Al^{3+}.

....

....

....

....

.... [3]

(b) Fig. 2.1 shows an incomplete energy cycle.

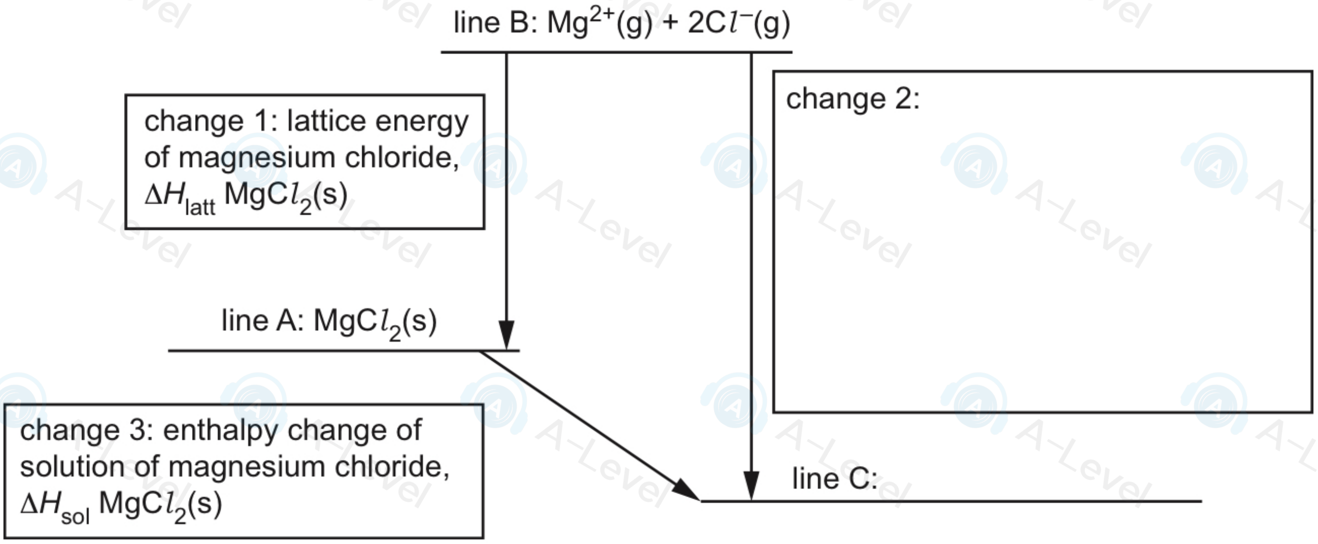


Fig. 2.1

(i) Complete line C on Fig. 2.1. Include state symbols. [1]

(ii) Use **both** words **and** symbols to identify change 2 on Fig. 2.1.

Use changes 1 and 3 as examples of how this should be done.

[2]

(iii) Calculate a value for the lattice energy of magnesium chloride, ΔH_{latt} $MgCl_2(s)$, by selecting and using appropriate data from Table 2.1.

Table 2.1

energy change	value / $kJ\,mol^{-1}$
enthalpy change of solution of magnesium chloride	–155
enthalpy change of formation of magnesium chloride	–642
first ionisation energy of magnesium	+736
second ionisation energy of magnesium	+1450
electron affinity of chlorine	–349
enthalpy change of hydration of Mg^{2+}	–1920
enthalpy change of hydration of Cl^-	–364

ΔH_{latt} $MgCl_2(s)$ = ... $kJ\,mol^{-1}$ [3]

(c) Define entropy.

....

.... [1]

(d) At 25 °C the enthalpy change of solution of compound **Z** is $+26\,kJ\,mol^{-1}$. The entropy change of solution of **Z** at the same temperature is $+52\,J\,K^{-1}\,mol^{-1}$.

Calculate the value of the Gibbs free energy change, ΔG, for the solution of **Z** at 25 °C.

ΔG = ... $kJ\,mol^{-1}$ [2]

(e) **(i)** Use your answer to **(d)** to predict whether or not **Z** is soluble in water at 25 °C. Explain your answer.

....

.... [1]

(ii) Predict whether **Z** becomes more or less soluble as the water is heated from 25 °C to 95 °C. Explain your answer.

....

.... [1]

[Total: 14]

7 **(a)** **(i)** Define the term *electron affinity*.

....

....

.... [2]

(ii) Define the term *lattice energy*.

....

....

.... [2]

(b) Use the following data and relevant data from the *Data Booklet* to calculate a value for the enthalpy change of formation of zinc bromide, $ZnBr_2(s)$.

You might find it helpful to construct an energy cycle.

electron affinity of Br(g) = $-325\,kJ\,mol^{-1}$
enthalpy change of atomisation of Zn(s) = $+131\,kJ\,mol^{-1}$
enthalpy change of vaporisation of $Br_2(l)$ = $+31\,kJ\,mol^{-1}$
lattice energy of $ZnBr_2(s)$ = $-2678\,kJ\,mol^{-1}$

enthalpy change of formation of $ZnBr_2(s)$ = ... $kJ\,mol^{-1}$ [4]

1 **(a)** The most common zinc mineral contains zinc(II) sulfide, ZnS.

(i) Complete the electrons in boxes diagram in Fig. 1.1 to show the electronic configuration of a zinc(II) ion.

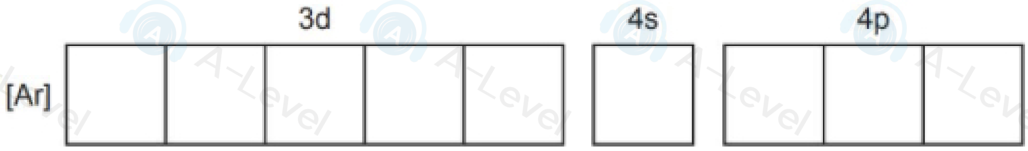


Fig. 1.1

[1]

(ii) Complete Fig. 1.2 to show the Born–Haber diagram for the ionic solid ZnS.

Include state symbols of relevant species.

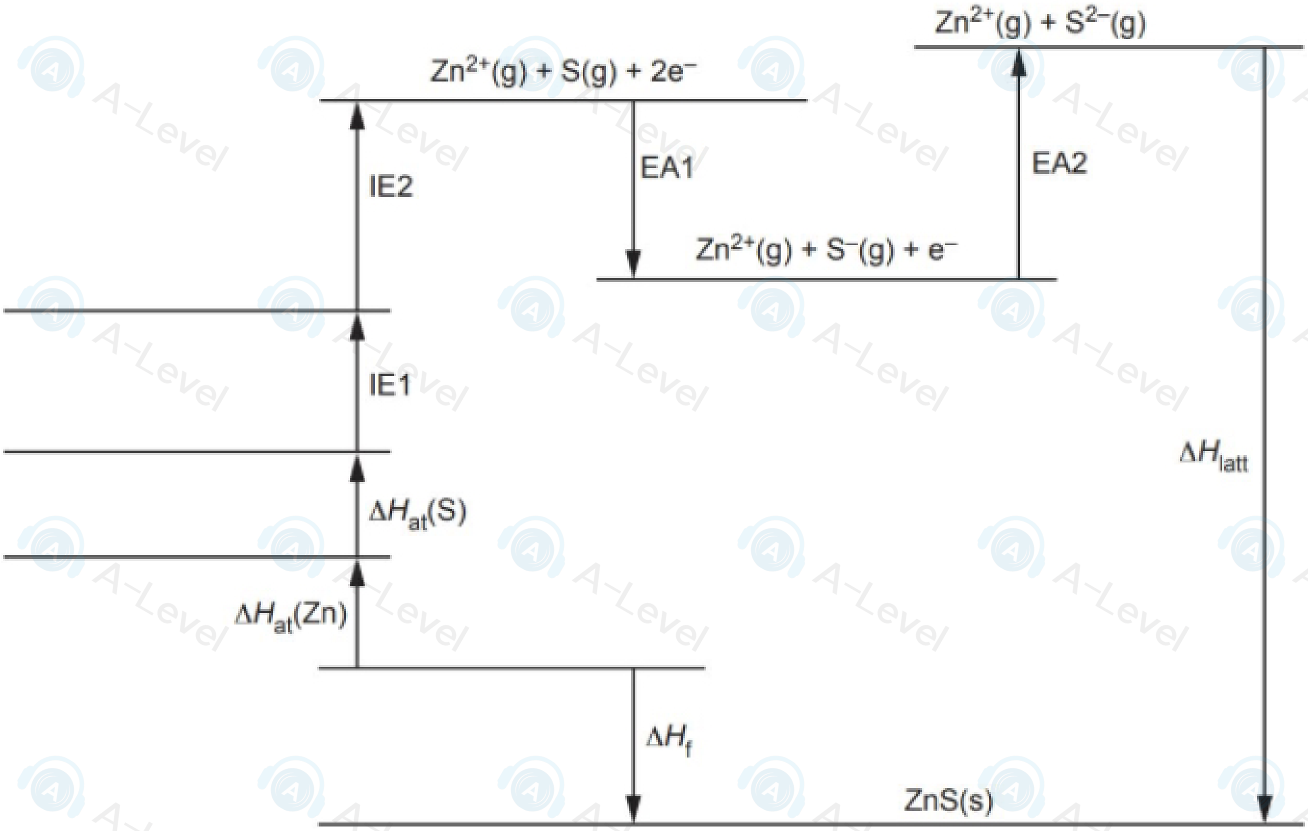


Fig. 1.2

[3]

(b) Zinc metal can be obtained in a two-step process as shown.

step 1 $2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$

step 2 $ZnO(s) + C(s) \rightarrow Zn(l) + CO(g)$

The reactions are carried out at 800 °C.

(i) Predict the sign of the entropy change, $\Delta S^{\ominus}$, of the reaction in **step 1**.

Explain your answer.

....

.... [1]

(ii) Use the data in Table 1.1 to calculate $\Delta S^{\ominus}$ of the reaction shown in **step 2**.

Table 1.1

chemical	ZnO(s)	C(s)	Zn(l)	CO(g)
$S^{\ominus}$ / $JK^{-1}mol^{-1}$	43.7	5.7	50.8	197.7

$\Delta S^{\ominus}$ = ... $JK^{-1}mol^{-1}$ [1]

(iii) An equation for the direct reduction of ZnS by carbon is shown.

$$2ZnS(s) + C(s) \rightarrow 2Zn(l) + CS_2(g) \qquad \Delta H^{\ominus} = +733\,kJ\,mol^{-1}$$
$$\Delta S^{\ominus} = +218\,J\,K^{-1}\,mol^{-1}$$

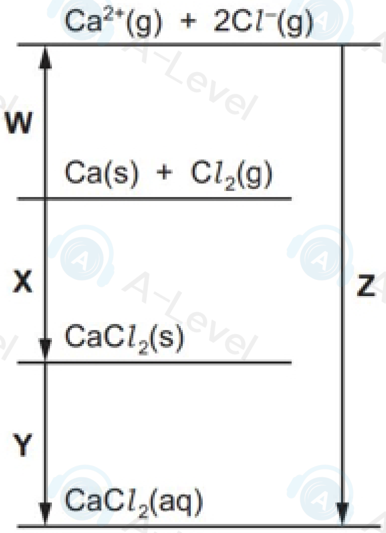
This reaction is **not** feasible at 800 °C.

Calculate $\Delta G^{\ominus}$ for this reaction at 800 °C.

$\Delta G^{\ominus}$ = ... $kJ\,mol^{-1}$ [2]

(c) $Zn(NO_3)_2$ undergoes thermal decomposition when heated. The reaction is similar to the thermal decomposition of Group 2 nitrates.

3 **(a)** The energy cycle shown can be used, along with suitable data, to calculate the enthalpy change of hydration of $Ca^{2+}(g)$.
Each arrow indicates a transformation, **W**, **X**, **Y** and **Z**. Each transformation consists of one or more steps.



The following data and data from the *Data Booklet* should be used.

electron affinity of $Cl(g)$ = –349 kJ mol^{-1}
enthalpy change of atomisation of $Ca(s)$ = +193 kJ mol^{-1}
enthalpy change of formation of $CaCl_2(s)$ = –795 kJ mol^{-1}
enthalpy change of solution of $CaCl_2(s)$ = –83 kJ mol^{-1}
enthalpy change of hydration of $Cl^{-}(g)$ = –364 kJ mol^{-1}

(i) Calculate the value of the enthalpy change corresponding to transformation **W**.
Show your working.

enthalpy change **W** = ... kJ mol^{-1} [2]

(ii) Use your answer to **(a)(i)** and other data to calculate the value of the enthalpy change corresponding to transformation **Z**.

enthalpy change **Z** = ... kJ mol^{-1} [2]

(iii) Use your answer to **(a)(ii)** to calculate the enthalpy change of hydration of $Ca^{2+}(g)$.

enthalpy change of hydration of $Ca^{2+}(g)$ = ... $kJ\,mol^{-1}$ [2]

(iv) Write an expression, in terms of **W**, **X**, **Y** and/or **Z**, to show how the enthalpy changes of **two** of the transformations can be used to calculate the lattice energy of $CaCl_2(s)$.

lattice energy of $CaCl_2(s)$ = ... [1]

(v) State whether the lattice energy of $CaCl_2(s)$ is more or less exothermic than the lattice energy of $MgF_2(s)$.

Explain your answer.

....

....

.... [1]

(b) The sulfates of the Group 2 elements vary in solubility down Group 2.

(i) Give the names of **two** solutions that could be mixed to form barium sulfate.

.... [1]

(ii) State and explain how the solubilities of the sulfates of the Group 2 elements vary down Group 2.

....

....

....

....

....

....

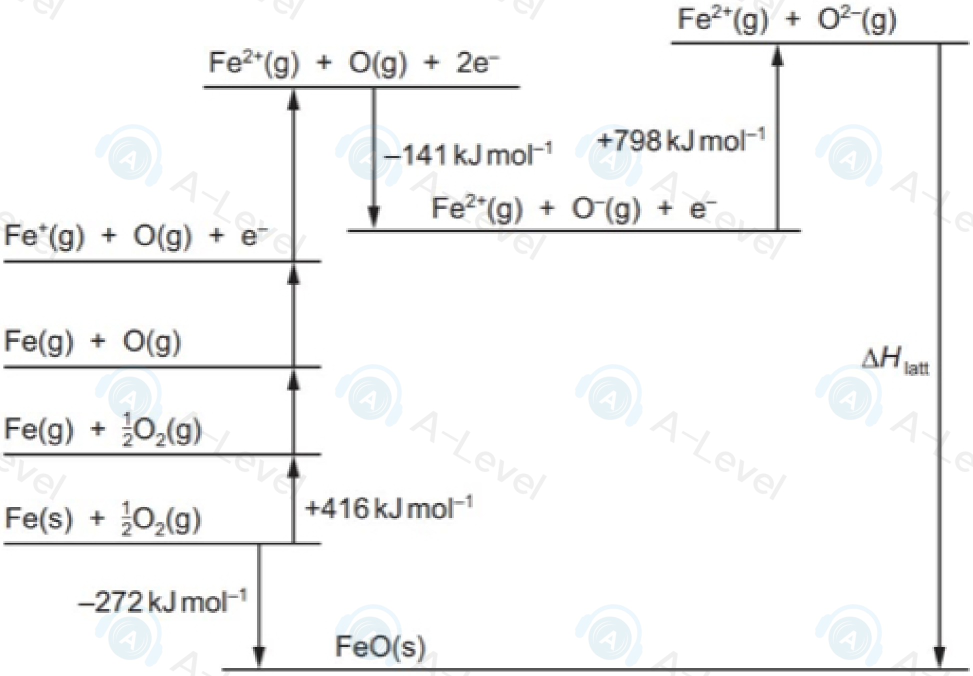
.... [4]

[Total: 13]

2 **(a)** Iron(II) compounds are generally only stable in neutral, non-oxidising conditions.

It is difficult to determine the lattice energy of FeO experimentally.

(i) Use data from the *Data Booklet* and this Born–Haber cycle to calculate the lattice energy ΔH_{latt}, of FeO(s) in kJ mol^{-1}.



ΔH_{latt}FeO(s) = ... kJ mol^{-1} [2]

(ii) Write the half-equation for the reaction that occurs at the anode during the electrolysis of $Fe_3O_4(l)$.

.... [1]

(iii) Calculate the maximum mass of iron metal formed when $Fe_3O_4(l)$ is electrolysed for six hours using a current of 50A.

Assume the one Fe^{2+} and two Fe^{3+} ions are discharged at the sam

6/28

mass of iron = ... g [3]

(c) $LiFePO_4$ can be used in lithium-ion rechargeable batteries.

When the cell is charging, lithium reacts with a graphite electrode to form LiC_6.

When the cell is discharging, the half-equations for the two processes that occur are as follows.

anode half-equation $LiC_6 \rightarrow 6C + Li^+ + e^-$

cathode half-equation $Li^+ + FePO_4 + e^- \rightarrow LiFePO_4$

(i) State one possible advantage of developing cells such as lithium-ion rechargeable

1 **(a)** Describe the trend in the solubility of the sulfates of magnesium, calcium and strontium.

Explain your answer.

.... > .... > ....

most soluble ... least soluble

....

....

....

....

....

....

[4]

(b) Define lattice energy, ΔH_{latt}.

....

.... [2]

(c) State and explain the main factors that affect the magnitude of lattice energies.

....

....

....

....

.... [2]

(d) Table 1.1 shows some energy changes.

Table 1.1

energy change	value / kJ mol^{-1}
standard enthalpy change of atomisation of potassium	+89
first ionisation energy of potassium	+419
second ionisation energy of potassium	+3070
standard enthalpy change of atomisation of sulfur	+279
S–S bond energy	+265
first ionisation energy of sulfur	+1000
second ionisation energy of sulfur	+2260
first electron affinity of sulfur	–200
second electron affinity of sulfur	+640
standard enthalpy change of formation of potassium sulfide, $K_2S(s)$	–381

(i) Born–Haber cycles can be used to determine the lattice energies of ionic compounds.

Complete the Born–Haber cycle in Fig. 1.1 for potassium sulfide, $K_2S(s)$.

Include state symbols for all of the species.

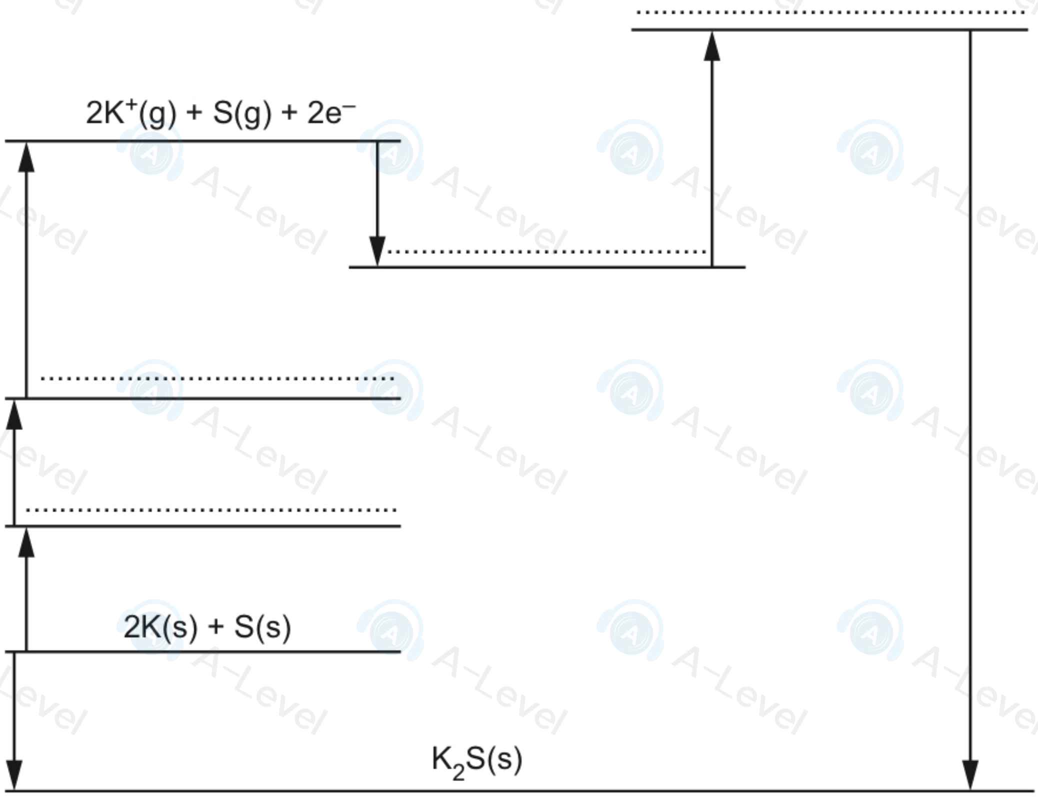


Fig. 1.1

[3]

(ii) Calculate the lattice energy, $\Delta H^\ominus_{latt}$, of $K_2S(s)$ using relevant data from Table 1.1.

Show your working.

$\Delta H^\ominus_{latt}$ of $K_2S(s)$ = ... kJ mol^{-1} [2]

[Total: 13]