

1 (a) Solutions of Group 2 hydrogencarbonates, $M(\text{HCO}_3)_2$, decompose on heating to give the corresponding metal carbonate, carbon dioxide and water.

(i) Write an equation for the decomposition of strontium hydrogencarbonate, $\text{Sr}(\text{HCO}_3)_2$.
..... [1]

(ii) The thermal stability of Group 2 carbonates increases down the group.
Explain this trend.
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.....
.....
..... [2]

(b) The hydroxides and fluorides of Group 2 elements show similar trends in solubility.

Describe the trend in the solubility of the fluorides of calcium, strontium and barium.

Explain your answer.

.....
least soluble most soluble
explanation
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.....
.....
.....
..... [4]

(c) (i) Define enthalpy change of hydration, ΔH_{hyd} .
.....
.....
..... [1]

(ii) State the main factors that affect the magnitude of enthalpy change of hydration.

Explain your answer.
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.....
.....
..... [2]

(d) Table 1.1 shows various energy changes.

Table 1.1

energy change	value / kJ mol^{-1}
lattice energy of MgF_2	–2957
enthalpy change of hydration, ΔH_{hyd} , of Mg^{2+}	–1926
enthalpy change of hydration, ΔH_{hyd} , of F^-	–505

Use data from Table 1.1 to calculate the enthalpy change of solution, ΔH_{sol} , for $\text{MgF}_2(\text{s})$.

It may be helpful to draw a labelled energy cycle. Show your working.

ΔH_{sol} of $\text{MgF}_2(\text{s}) = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

(e) Mercury(I) fluoride, Hg_2F_2 , is sparingly soluble in water.

The cation in Hg_2F_2 exists as the diatomic ion Hg_2^{2+} with a covalent Hg–Hg bond.

(i) Write the expression for the solubility product, K_{sp} , of Hg_2F_2 . Include the units.

$$K_{\text{sp}} =$$

units [2]

(ii) The solubility of Hg_2F_2 is $9.20 \times 10^{-3} \text{ mol dm}^{-3}$ at 298 K.

Calculate the value of K_{sp} of Hg_2F_2 at 298 K.

$$K_{\text{sp}} = \dots\dots\dots [1]$$

[Total: 15]

4 (a) Define enthalpy change of atomisation, ΔH_{at} .

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..... [1]

(b) Define first electron affinity, EA.

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..... [1]

(c) Explain why the first electron affinity of chlorine is more exothermic than the first electron affinity of iodine.

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.....
..... [2]

(d) The enthalpy change for the reaction $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{g})$ is -486 kJ mol^{-1} .

The first electron affinity of chlorine is -364 kJ mol^{-1} .

Calculate the enthalpy change of atomisation of chlorine.

$$\Delta H_{\text{at}} \text{ of chlorine} = \dots\dots\dots \text{ kJ mol}^{-1} [2]$$

[Total: 6]

- 2 (a) The Group 2 sulfates and the Group 2 chromates show similar trends in solubility.

Suggest the trend in the solubility of the Group 2 chromates down the group.

Explain your answer.

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..... [4]

- (b) Silver(I) chromate, Ag_2CrO_4 , is sparingly soluble in water.

- (i) Write an ionic equation to show the equilibrium between solid Ag_2CrO_4 and its aqueous solution.

Include state symbols.

..... [1]

- (ii) The value of the solubility product, K_{sp} , of Ag_2CrO_4 is 1.12×10^{-12} at 298 K.

Calculate the equilibrium concentration of Ag^+ , in mol dm^{-3} , in a saturated solution of Ag_2CrO_4 at 298 K.

equilibrium concentration of Ag^+ = mol dm^{-3} [3]

- (c) The hydrogenchromate ion, HCrO_4^- , is a weak acid. The $\text{p}K_{\text{a}}$ of HCrO_4^- is 6.49.

- (i) Calculate the pH of a $0.0250 \text{ mol dm}^{-3}$ HCrO_4^- solution.

pH = [2]

- (ii) HCrO_4^- can show amphoteric behaviour.

State the formula of:

- the conjugate acid of HCrO_4^-
- the conjugate base of HCrO_4^- [1]

(d) Table 2.1 shows some energy changes.

Table 2.1

energy change	value/kJ mol ⁻¹
first ionisation energy of silver	+731
second ionisation energy of silver	+2074
first ionisation energy of sulfur	+1000
second ionisation energy of sulfur	+2251
first electron affinity of sulfur	-200
second electron affinity of sulfur	+532
enthalpy change of atomisation of sulfur	+279
enthalpy change of formation of silver(I) sulfide, Ag ₂ S(s)	-33
lattice energy of silver(I) sulfide, Ag ₂ S(s)	-2677

(i) Define the term first electron affinity.

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..... [1]

(ii) Explain why the value for the second electron affinity of sulfur is positive.

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.....
..... [1]

(iii) Construct an equation for the lattice energy of Ag₂S.

Include state symbols.

..... [1]

(iv) Calculate the enthalpy change of atomisation, ΔH_{at} , in kJ mol⁻¹, of silver using relevant data from Table 2.1.

It may be helpful to draw a labelled Born–Haber cycle.

Show your working.

ΔH_{at} of silver = kJ mol⁻¹ [3]

(e) Suggest how the magnitude for the lattice energy of Ag₂S(s) differs from the lattice energy of Cu₂S(s).

Explain your answer.

.....
.....
..... [1]

[Total: 18]

3 (a) Define the term entropy.

.....
 [1]

(b) (i) Place **one** tick (✓) in each row of Table 3.1 to show the sign of the entropy change, ΔS , for each process.

Table 3.1

process	ΔS is negative	ΔS is positive
steam condensing into water		
solid KCl dissolving in water		

[1]

(ii) Chlorine trifluoride, ClF_3 , decomposes on heating into its elements, as shown.



Standard entropies are shown in Table 3.2.

Table 3.2

substance	$\text{ClF}_3(\text{g})$	$\text{Cl}_2(\text{g})$	$\text{F}_2(\text{g})$
$S^\ominus / \text{JK}^{-1} \text{mol}^{-1}$	+281.6	+223.1	+203.0

Calculate the standard entropy change, $\Delta S^\ominus$, in $\text{JK}^{-1} \text{mol}^{-1}$, for reaction 1.

$\Delta S^\ominus$ for reaction 1 = $\text{JK}^{-1} \text{mol}^{-1}$ [2]

(c) Group 2 carbonates decompose on heating. The decomposition for one of the Group 2 carbonates, MCO_3 , is shown in reaction 2.



(i) Predict the sign of the entropy change, ΔS , for reaction 2.

Explain your answer.

.....
 [1]

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(ii) The Gibbs equation is shown.

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

Fig. 3.1 shows values of the Gibbs free energy change, $\Delta G^\ominus$, in kJ mol^{-1} , at different temperatures, T , in K, for reaction 2.

Assume $\Delta H^\ominus$ and $\Delta S^\ominus$ values for this reaction remain constant over this temperature range.

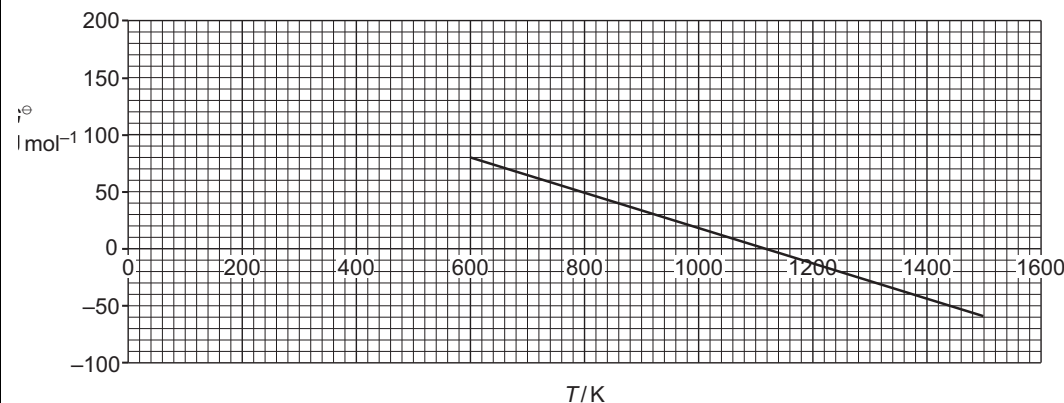


Fig. 3.1

Use the gradient and intercept on the y-axis in Fig. 3.1 and the Gibbs equation to determine:

- $\Delta S^\ominus$, in $\text{JK}^{-1} \text{mol}^{-1}$, for reaction 2
- the minimum temperature, T , in K, at which the reaction is feasible
- $\Delta H^\ominus$, in kJ mol^{-1} , for reaction 2.

$\Delta S^\ominus$ for reaction 2 = $\text{JK}^{-1} \text{mol}^{-1}$

minimum temperature, T = K

$\Delta H^\ominus$ for reaction 2 = kJ mol^{-1}

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