1. 9701/41/O/N/16 Q3

The spontaneity (feasibility) of a chemical reaction depends on the standard Gibbs free energy change, $\Delta G^{\ominus}$. This is related to the standard enthalpy and entropy changes by the equation shown.

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

(a) State and explain whether the following processes will lead to an increase or decrease in entropy.

(i) the reaction of magnesium with hydrochloric acid

entropy change ....

explanation ....
[1]

(ii) solid potassium chloride dissolving in water

entropy change ....

explanation ....
[1]

(iii) steam condensing to water

entropy change ....

explanation ....
[1]

(b) Magnesium carbonate can be decomposed.

$$MgCO_3(s) \rightarrow MgO(s) + CO_2(g) \qquad \Delta H^{\ominus} = +117\,kJ\,mol^{-1}$$

Standard entropies are shown in the table.

substance	$MgCO_3(s)$	$MgO(s)$	$CO_2(g)$
$S^{\ominus}$ / J mol^{-1} K^{-1}	+65.7	+26.9	+214

(i) Calculate $\Delta G^{\ominus}$ for this reaction at 298 K.
Include a relevant sign and give your answer to **three** significant figures.

$\Delta G^{\ominus}$ = ... kJ mol^{-1} [3]

(ii) Explain, with reference to $\Delta G^{\ominus}$, why this reaction becomes more feasible at higher temperatures.

....

.... [1]

(c) On heating, sodium hydrogencarbonate decomposes into sodium carbonate as shown.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g) \qquad \Delta H^\ominus = +130\,kJ\,mol^{-1}$$
$$\Delta S^\ominus = +316\,J\,mol^{-1}\,K^{-1}$$

Calculate the **minimum** temperature at which this reaction becomes spontaneous (feasible).
Show your working.

temperature = ... K [2]

(d) The solubility of Group 2 sulfates decreases down the Group.

Explain this trend.

....

....

....

.... [2]

[Total: 11]

2. 9701/42/O/N/16 Q2c(iii)

(iii) Use the following data and any relevant data from the *Data Booklet* to calculate the standard enthalpy change of formation, $\Delta H_f^\ominus$, of $NaN_3(s)$.
Include a sign in your answer. Show all your working.

lattice energy, $\Delta H_{latt}^\ominus$, of $NaN_3(s)$	$-732\,kJ\,mol^{-1}$
standard enthalpy change of atomisation, $\Delta H_{at}^\ominus$, of Na(g)	$+107\,kJ\,mol^{-1}$
standard enthalpy change, $\Delta H^\ominus$, for $1\frac{1}{2}N_2(g) + e^- \rightarrow N_3^-(g)$	$+142\,kJ\,mol^{-1}$

$\Delta H_f^\ominus$ of $NaN_3(s)$ = ... $kJ\,mol^{-1}$ [3]

(iv) The lattice energy, $\Delta H_{latt}^\ominus$, of $RbN_3(s)$ is $-636\,kJ\,mol^{-1}$.

Suggest why the lattice energy of $NaN_3(s)$, $-732\,kJ\,mol^{-1}$, is more exothermic than that of $RbN_3(s)$.

....

.... [1]

[Total: 11]

3. 9701/42/O/N/16 3c

(c) Iron(III) oxide can be reduced to iron metal using carbon monoxide at a temperature of 1000 °C.

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g) \qquad \Delta H^{\ominus} = -43.6\,kJ\,mol^{-1}$$

Some relevant standard entropies are given in the table.

substance	$Fe_2O_3(s)$	$CO(g)$	$Fe(s)$	$CO_2(g)$
$S^{\ominus}$ / J K^{-1} mol^{-1}	+90	+198	+27	+214

(i) What is meant by the term *entropy*?

....

.... [1]

(ii) Calculate the standard entropy change, $\Delta S^{\ominus}$, for this reaction.

$\Delta S^{\ominus}$ = ... J K^{-1} mol^{-1} [2]

(iii) Calculate the standard Gibbs free energy change, $\Delta G^{\ominus}$, for this reaction at 25 °C.

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

(iv) Suggest why a temperature of 1000 °C is usually used for this reaction, even though the reaction is spontaneous (feasible) at 25 °C. Explain your answer.

....

....

.... [1]

[Total: 13]

4. 9701/41/M/J/16 Q8

(a) Chromyl chloride, CrO_2Cl_2, can be prepared by heating a mixture of potassium dichromate(VI) and potassium chloride with concentrated sulfuric acid.

$$K_2Cr_2O_7 + 4KCl + 3H_2SO_4 \rightarrow 2CrO_2Cl_2 + 3K_2SO_4 + 3H_2O$$

Use the following data to complete the Hess' Law cycle and calculate the enthalpy change of the reaction, ΔH_r.

compound	enthalpy change of formation, $\Delta H_f^\ominus$ / kJ mol^{-1}
$K_2Cr_2O_7$	–2061
KCl	–437
H_2SO_4	–814
CrO_2Cl_2	–580
K_2SO_4	–1438
H_2O	–286

$$\boxed{K_2Cr_2O_7 + 4KCl + 3H_2SO_4} \xrightarrow{\Delta H_r} \boxed{2CrO_2Cl_2 + 3K_2SO_4 + 3H_2O}$$

$$\boxed{\text{elements}}$$

ΔH_r = ... kJ mol^{-1} [2]

5. 9701/42/M/J/16 Q5b

(b) Values for $\Delta H^{\ominus}$ and $\Delta G^{\ominus}$ for equilibria **I** and **II**, and the value of $\Delta S^{\ominus}$ for equilibrium **I**, are given in the table below. All values are at a temperature of 298 K.

equilibrium	$\Delta H^{\ominus}$ / kJ mol^{-1}	$\Delta G^{\ominus}$ / kJ mol^{-1}	$\Delta S^{\ominus}$ / J K^{-1} mol^{-1}
I	–57.3	–37.4	–66.8
II	–56.5	–60.7	to be calculated

(i) Suggest a reason why the $\Delta H^{\ominus}$ values for the two equilibria are very similar.

....

.... [1]

(ii) Calculate $\Delta S^{\ominus}$ for equilibrium **II**.

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

(iii) Suggest a reason for the difference between the $\Delta S^{\ominus}$ you have calculated for equilibrium **II** and that for equilibrium **I** given in the table.

....

....

.... [1]

(iv) Which of the two complexes is the more stable? Give a reason for your answer.

....

.... [1]

[Total: 8]

6. 9701/42/M/J/16 Q8(ii)

(ii) The ionic equation for this reaction is shown.

$$Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g) \qquad \Delta H^\ominus = \mathbf{x}\ kJ\,mol^{-1}$$

Construct a **fully labelled** Hess' Law cycle to connect each side of this equation to the relevant gas phase ions.

Use your cycle, the following data, **and** data from the *Data Booklet*, to calculate a value for **x**.

standard enthalpy of atomisation of Ca(s), $\Delta H^\ominus_{at}(Ca)$	+178 kJ mol^{-1}
standard enthalpy of hydration of Ca^{2+}(g), $\Delta H^\ominus_{hyd}(Ca^{2+})$	–1576 kJ mol^{-1}
standard enthalpy of hydration of H^+(g), $\Delta H^\ominus_{hyd}(H^+)$	–1090 kJ mol^{-1}

x = ... kJ mol^{-1} [4]

(c) The standard enthalpy change for the reaction between Ca(s) and CH_3CO_2H(aq) is **less negative** than **x** by 2 kJ mol^{-1}.

Suggest an explanation for this.

....

.... [2]

[Total: 10]

7. 9701/43/M/J/16 Q3(c)

(c) **(i)** Predict the sign of $\Delta S^{\ominus}$ for this reaction. Explain your answer.

$$2H_2S(g) + CH_4(g) \rightleftharpoons CS_2(g) + 4H_2(g) \qquad \Delta H^{\ominus} = +241\,\text{kJ mol}^{-1}$$

....

.... [1]

The free energy change, $\Delta G^{\ominus}$, for this reaction at 1000 K is +51 kJ mol^{-1}.

(ii) Calculate the value of $\Delta S^{\ominus}$ for this reaction, stating its units.

$\Delta S^{\ominus}$ = ... units .... [2]

(d) How would the value of $\Delta G^{\ominus}$, and hence the spontaneity (feasibility) of this reaction change as the temperature increases? Explain your answer.

....

....

.... [2]

[Total: 10]

8. 9701/43/M/J/16 Q8

(a) Chromyl chloride, CrO_2Cl_2, can be prepared by heating a mixture of potassium dichromate(VI) and potassium chloride with concentrated sulfuric acid.

$$K_2Cr_2O_7 + 4KCl + 3H_2SO_4 \rightarrow 2CrO_2Cl_2 + 3K_2SO_4 + 3H_2O$$

Use the following data to complete the Hess' Law cycle and calculate the enthalpy change of the reaction, ΔH_r.

compound	enthalpy change of formation, $\Delta H_f^\ominus$ / kJ mol^{-1}
$K_2Cr_2O_7$	–2061
KCl	–437
H_2SO_4	–814
CrO_2Cl_2	–580
K_2SO_4	–1438
H_2O	–286

$$\boxed{K_2Cr_2O_7 + 4KCl + 3H_2SO_4} \xrightarrow{\Delta H_r} \boxed{2CrO_2Cl_2 + 3K_2SO_4 + 3H_2O}$$

$$\boxed{\text{elements}}$$

ΔH_r = ... kJ mol^{-1} [2]

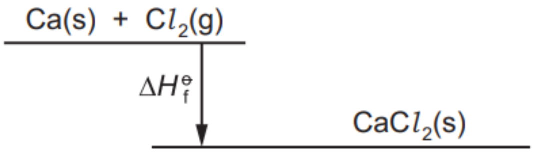
9. 9701/42/F/M/16 Q2

(a) Calcium metal reacts with chlorine gas to form calcium chloride, $CaCl_2$.

(i) Write an equation, including state symbols, to represent the lattice energy of calcium chloride, $CaCl_2$.

.... [1]

(ii) Complete a fully labelled Born-Haber cycle that could be used to calculate the lattice energy, $\Delta H^{\ominus}_{latt}$, for calcium chloride.



[2]

(iii) Use your answer to **(ii)** and the following data, together with relevant data from the *Data Booklet*, to calculate a value for $\Delta H^{\ominus}_{latt}$ for calcium chloride.

standard enthalpy change of formation of $CaCl_2(s)$, $\Delta H^{\ominus}_{f}$	$-796\,kJ\,mol^{-1}$
standard enthalpy change of atomisation of Ca(s), $\Delta H^{\ominus}_{at}$	$+178\,kJ\,mol^{-1}$
electron affinity of chlorine atoms	$-349\,kJ\,mol^{-1}$

$\Delta H^{\ominus}_{latt}$ = ... $kJ\,mol^{-1}$ [3]

(b) Entropy is a measure of the disorder of a system.

Describe and explain what happens to the entropy of a gas when the temperature is increased.

....

....

.... [2]

(c) The table shows four reactions.

(i) For each reaction, predict the sign of the entropy change, $\Delta S^\ominus$. If you predict no entropy change, write 'no change' in the table below. The first one has been done for you.

reaction	sign of $\Delta S^\ominus$
$CO(g) + O_2(g) \rightarrow CO_2(g)$	negative
$Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$	
$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4.5H_5O(s)$	
$NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + CO_2(g) + H_2O(l)$	

[2]

(ii) Explain why the entropy change for the first process is negative.

....

.... [1]

(d) Calculate the standard entropy change, $\Delta S^\ominus$, for this reaction.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Standard entropies, $S^\ominus$, in $J\,K^{-1}\,mol^{-1}$ are given.

$N_2(g)$	$H_2(g)$	$NH_3(g)$
+192	+131	+193

$\Delta S^\ominus$ .... $J\,K^{-1}\,mol^{-1}$ [2]

(e) Whether or not a chemical reaction is spontaneous (feasible) can be deduced by calculating the change in free energy, $\Delta G^\ominus$, at a given temperature.

$$MgCO_3(s) \rightarrow MgO(s) + CO_2(g) \qquad \Delta H^\ominus = +117\,kJ\,mol^{-1}$$
$$\Delta S^\ominus = +175\,J\,K^{-1}\,mol^{-1}$$

(i) Calculate the value of $\Delta G^\ominus$ at 298 K for the above reaction.

[2]

(ii) Use your answer to **(i)** to explain whether or not this reaction is spontaneous at 298 K.

....

.... [1]

[Total: 16]

10. 9701/42/O/N/17 Q5

(a) The arrangement of the anions around a cation is called the geometry of the cation; e.g. in $[CuCl_4]^{2-}$ the geometry of copper is tetrahedral and the co-ordination number of copper is 4.

The geometry of a cation in an ionic compound can be predicted from the ratio of the ionic radii of the cation and anion involved.

$\frac{\text{cation radius}}{\text{anion radius}}$	geometry of cation
0.155–0.225	trigonal planar
0.225–0.414	tetrahedral
0.414–0.732	octahedral

Use data from the *Data Booklet* to predict the geometry of, and hence the co-ordination number of, the cation for

- sodium chloride, NaC*l*,

geometry of Na^+ = ... co-ordination number of Na^+ = ...

- magnesium chloride, $MgCl_2$.

geometry of Mg^{2+} = ... co-ordination number of Mg^{2+} = ...

[2]

(b) Magnesium(I) chloride, MgC*l*, is an unstable compound and readily decomposes as shown.

$$2MgCl(s) \rightarrow Mg(s) + MgCl_2(s)$$

Use the following data to calculate the enthalpy change of this reaction.

$\Delta H_f^\ominus$ $MgCl(s)$ = –106 kJ mol^{-1}
$\Delta H_f^\ominus$ $MgCl_2(s)$ = –642 kJ mol^{-1}

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

(c) **(i)** The equation for which ΔH is the lattice energy for MgC*l* is shown.

$$Mg^+(g) + Cl^-(g) \rightarrow MgCl(s)$$

Use the equation, the following data, and relevant data from the *Data Booklet* to calculate a value for the lattice energy of MgC*l*. You might find it helpful to construct an energy cycle.

electron affinity of C*l*(g) = –349 kJ mol^{-1}
enthalpy change of atomisation of Mg(s) = +147 kJ mol^{-1}
enthalpy change of formation of MgC*l*(s) = –106 kJ mol^{-1}

lattice energy MgC*l* = ... kJ mol^{-1} [3]

(ii) Suggest how the lattice energies of $MgCl_2$ and NaC*l* will compare to that of MgC*l*. Explain your answers.

$MgCl_2$ and MgC*l* ....

....

NaC*l* and MgC*l*..

....

[3]

(d) Define the term *electron affinity*.

....

.... [2]

[Total: 11]

11. 9701/41/M/J/17 Q1

The table lists the standard enthalpy changes of formation, $\Delta H_f^\ominus$, for some compounds and aqueous ions.

species	$\Delta H_f^\ominus$ / kJ mol^{-1}
$Ba^{2+}(aq)$	–538
$OH^-(aq)$	–230
$CO_2(g)$	–394
$BaCO_3(s)$	–1216
$H_2O(l)$	–286

(b) **(i)** Reaction 1 occurs when $CO_2(g)$ is bubbled through an aqueous solution of $Ba(OH)_2$.

Use the data in the table to calculate the standard enthalpy change for reaction 1, $\Delta H_{r1}^\ominus$.

$$Ba(OH)_2(aq) + CO_2(g) \rightarrow BaCO_3(s) + H_2O(l) \qquad \text{reaction 1}$$

$\Delta H_{r1}^\ominus$ = ... kJ mol^{-1} [2]

If $CO_2(g)$ is bubbled through an aqueous solution of $Ba(OH)_2$ for a long time, the precipitated $BaCO_3(s)$ dissolves, as shown in reaction 2.

$$BaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ba(HCO_3)_2(aq) \qquad \text{reaction 2}$$

The standard enthalpy change for reaction 2, $\Delta H^{\ominus}_{r2}$, = –26 kJ mol⁻¹.

(ii) Use this information and the data in the table to calculate the standard enthalpy change of formation of the $HCO_3^-(aq)$ ion.

$\Delta H^{\ominus}_{f}$ $HCO_3^-(aq)$ = ... kJ mol⁻¹ [2]

(iii) The overall process is shown by reaction 3.

Use your answer to **(ii)**, and the data given in the table, to calculate the standard enthalpy change for reaction 3, $\Delta H^{\ominus}_{r3}$.

$$Ba(OH)_2(aq) + 2CO_2(g) \rightarrow Ba(HCO_3)_2(aq) \qquad \text{reaction 3}$$

$\Delta H^{\ominus}_{r3}$ = ... kJ mol⁻¹ [1]

(iv) How would the value of $\Delta H^{\ominus}_{r3}$ compare with the value of $\Delta H^{\ominus}_{r4}$ for the similar reaction with $Ca(OH)_2(aq)$ as shown in reaction 4?
Explain your answer.

$$Ca(OH)_2(aq) + 2CO_2(g) \rightarrow Ca(HCO_3)_2(aq) \qquad \text{reaction 4}$$

....

....

.... [2]

(c) The standard entropy change for reaction 1 is $\Delta S^{\ominus}_{r1}$.

Suggest, with a reason, how the standard entropy change for reaction 3 might compare with $\Delta S^{\ominus}_{r1}$.

....

....

.... [2]

[Total: 13]

12. 9701/42/M/J/17 Q1

(c) **(i)** Use the data in the table below, and relevant data from the *Data Booklet*, to calculate the lattice energy, $\Delta H^{\ominus}_{latt}$, of potassium oxide, $K_2O(s)$.

energy change	value / kJ mol^{-1}
enthalpy change of atomisation of potassium, $\Delta H^{\ominus}_{at}$ K(s)	+89
electron affinity of O(g)	–141
electron affinity of O^-(g)	+798
enthalpy change of formation of potassium oxide, $\Delta H^{\ominus}_{f}$ $K_2O(s)$	–361

$\Delta H^{\ominus}_{latt}$ = ... kJ mol^{-1} [3]

(ii) State whether the lattice energy of Na_2O would be more negative, less negative or the same as that of K_2O. Give reasons for your answer.

....

.... [1]

[Total: 10]

13. 9701/43/M/J/17 Q1

The table lists the standard enthalpy changes of formation, $\Delta H_f^{\ominus}$, for some compounds and aqueous ions.

species	$\Delta H_f^{\ominus}$ / kJ mol^{-1}
$Ba^{2+}(aq)$	–538
$OH^-(aq)$	–230
$CO_2(g)$	–394
$BaCO_3(s)$	–1216
$H_2O(l)$	–286

(b) **(i)** Reaction 1 occurs when $CO_2(g)$ is bubbled through an aqueous solution of $Ba(OH)_2$.

Use the data in the table to calculate the standard enthalpy change for reaction 1, $\Delta H_{r1}^{\ominus}$.

$$Ba(OH)_2(aq) + CO_2(g) \rightarrow BaCO_3(s) + H_2O(l) \qquad \text{reaction 1}$$

$\Delta H_{r1}^{\ominus}$ = ... kJ mol^{-1} [2]

If $CO_2(g)$ is bubbled through an aqueous solution of $Ba(OH)_2$ for a long time, the precipitated $BaCO_3(s)$ dissolves, as shown in reaction 2.

$$BaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ba(HCO_3)_2(aq) \qquad \text{reaction 2}$$

The standard enthalpy change for reaction 2, ΔH°_{r2}, = –26 kJ mol^{-1}.

(ii) Use this information and the data in the table to calculate the standard enthalpy change of formation of the $HCO_3^-(aq)$ ion.

ΔH°_{f} $HCO_3^-(aq)$ = ... kJ mol^{-1} [2]

(iii) The overall process is shown by reaction 3.

Use your answer to **(ii)**, and the data given in the table, to calculate the standard enthalpy change for reaction 3, ΔH°_{r3}.

$$Ba(OH)_2(aq) + 2CO_2(g) \rightarrow Ba(HCO_3)_2(aq) \qquad \text{reaction 3}$$

ΔH°_{r3} = ... kJ mol^{-1} [1]

(iv) How would the value of ΔH°_{r3} compare with the value of ΔH°_{r4} for the similar reaction with $Ca(OH)_2(aq)$ as shown in reaction 4?
Explain your answer.

$$Ca(OH)_2(aq) + 2CO_2(g) \rightarrow Ca(HCO_3)_2(aq) \qquad \text{reaction 4}$$

....

....

.... [2]

(c) The standard entropy change for reaction 1 is ΔS°_{r1}.

Suggest, with a reason, how the standard entropy change for reaction 3 might compare with ΔS°_{r1}.

....

....

.... [2]

[Total: 13]

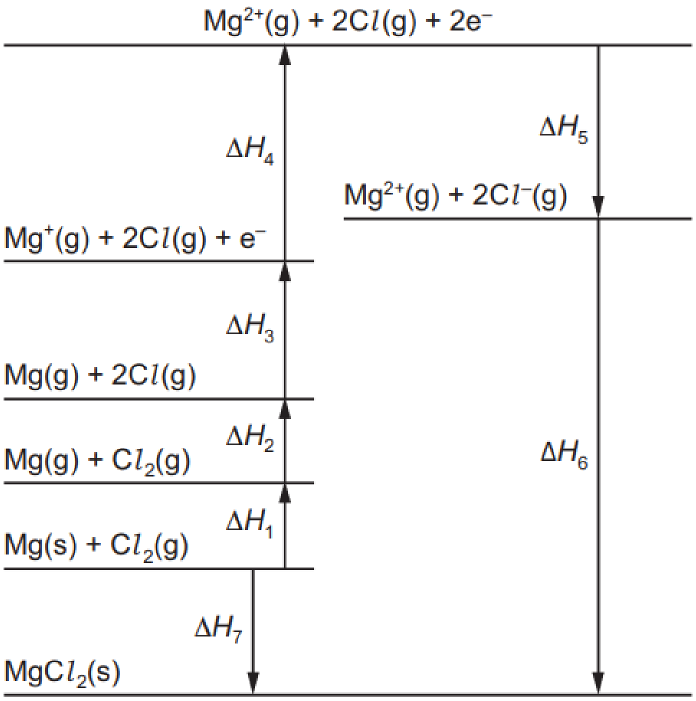
14. 9701/42/F/M/17 Q2

(a) Complete the table using ticks (✓) to indicate whether the sign of each type of energy change, under standard conditions, is always positive, always negative or could be either positive or negative.

energy change	always positive	always negative	either positive or negative
electron affinity			
enthalpy change of atomisation			
ionisation energy			
lattice energy			

[2]

(b) The Born-Haber cycle for magnesium chloride is shown.



(i) Explain why ΔH_4 is greater than ΔH_3.

....

.... [1]

(ii) What names are given to the enthalpy changes ΔH_6 and ΔH_7?

ΔH_6 ....

ΔH_7 ....

[1]

(c) Chlorine is in Group 17.

Suggest the trend in the first electron affinity of the elements in Group 17. Explain your answer.

....

....

.... [2]

(d) The equation for the formation of magnesium oxide from its elements is shown.

$$Mg(s) + \tfrac{1}{2}O_2(g) \rightarrow MgO(s) \qquad \Delta H^{\ominus} = -602\,kJ\,mol^{-1}$$

substance	$S^{\ominus}$ / J K^{-1} mol^{-1}
Mg(s)	32.7
O_2(g)	205
MgO(s)	26.9

Use the equation and the data given in the table to calculate $\Delta G^{\ominus}$ for the reaction at 25 °C.

$\Delta G^{\ominus}$ = ... units ....

[4]

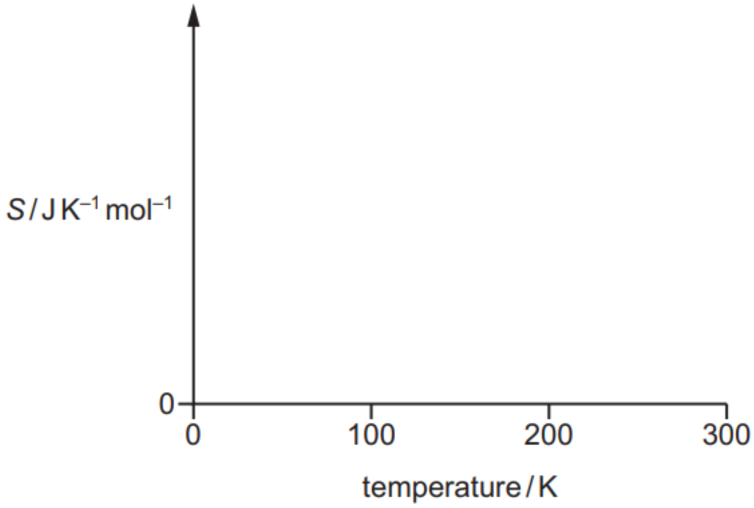
[Total: 10]

15. 9701/41/O/N/18 Q8

Entropy is a measure of the disorder of a system.

(a) Assume the entropy, *S*, for H_2O is zero at 0K.

Sketch a graph on the axes to show how the entropy changes for H_2O between 0K and 300K.



[2]

(b) Place **one tick** (✓) in **each row** of the table to show the sign of the entropy changes, Δ*S*.

	Δ*S* is negative	Δ*S* is positive
solid dissolving in water		
water boiling to steam		

[1]

(c) The equation for a reaction that produces methanol is shown.

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

Use relevant bond energies from the *Data Booklet* to calculate the enthalpy change, Δ*H*, for this gas phase reaction.

Δ*H* = ... kJ mol^{-1} [2]

(d) At 298 K, both products of this reaction are liquid.

$$CO_2(g) + 3H_2(g) \rightarrow CH_3OH(l) + H_2O(l) \qquad \Delta H^\ominus = -131\,kJ\,mol^{-1}$$

Standard entropies are shown in the table.

substance	$CO_2(g)$	$H_2(g)$	$CH_3OH(l)$	$H_2O(l)$
$S^\ominus / J\,K^{-1}\,mol^{-1}$	+214	+131	+127	+70

(i) Calculate the standard entropy change, $\Delta S^\ominus$, for this reaction.

$\Delta S^\ominus$ = ... $J\,K^{-1}\,mol^{-1}$ [2]

(ii) Calculate the standard Gibbs free energy change, $\Delta G^\ominus$, for this reaction at 298 K.

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

(iii) Predict the effect of increasing the temperature on the feasibility of this reaction.

....

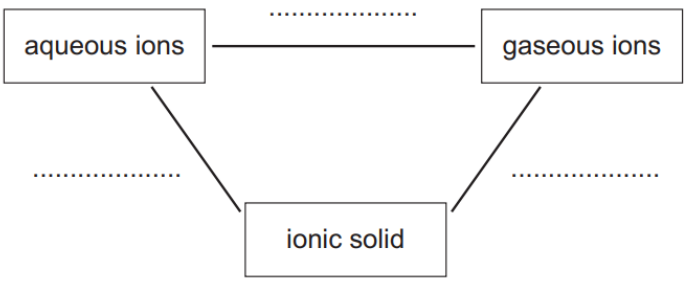
.... [1]

16. 9701/42/O/N/18 4c

(c) Complete the energy cycle to show the enthalpy changes that occur in the transformations between aqueous ions, gaseous ions and an ionic solid.

On your diagram label each enthalpy change with its appropriate symbol; lattice energy, $\Delta H^{\ominus}_{\text{latt}}$, enthalpy change of hydration, $\Delta H^{\ominus}_{\text{hyd}}$, or enthalpy change of solution, $\Delta H^{\ominus}_{\text{sol}}$.

Complete the three arrows showing the correct direction of each enthalpy change.



[3]

17. 9701/42/O/N/18 8c

(c) Some values for standard enthalpy changes of formation, $\Delta H^{\ominus}_{f}$, and standard entropies, $S^{\ominus}$, are given in the table.

	$\Delta H^{\ominus}_{f}$ / kJ mol^{-1}	$S^{\ominus}$ / J K^{-1} mol^{-1}
$NH_2Cl(g)$	+80.1	+241
$NH_3(g)$	–45.9	+198
$N_2H_4(g)$	+95.4	+237
$HCl(g)$	–92.3	+187

(i) Define the meaning of the term *entropy*.

....

.... [1]

Hydrazine, N_2H_4, can be produced from chloramine and ammonia as shown.

$$NH_2Cl(g) + NH_3(g) \rightarrow N_2H_4(g) + HCl(g)$$

(ii) Calculate the standard entropy change, $\Delta S^{\ominus}$, for this reaction.

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

(iii) Calculate the standard enthalpy change, $\Delta H^\ominus$, for this reaction.

$\Delta H^\ominus$ = ... kJ mol^{-1} [1]

(iv) Calculate the standard Gibbs free energy change, $\Delta G^\ominus$, for this reaction at 298 K.

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

(v) Explain, with reference to $\Delta G^\ominus$, why this reaction becomes **less** feasible at higher temperatures.

....

....

.... [1]

18. 9701/41/M/J/18 Q1(d)

(d) Use the data below, and other suitable data from the *Data Booklet*, to calculate the lattice energy of sodium oxide, $\Delta H^{\ominus}_{\text{latt}}$ $Na_2O(s)$.

energy change	value / kJ mol^{-1}
standard enthalpy change of formation of sodium oxide, $\Delta H^{\ominus}_{f}$ $Na_2O(s)$	–416
standard enthalpy change of atomisation of sodium, $\Delta H^{\ominus}_{at}$ Na(s)	+109
electron affinity of O(g)	–142
electron affinity of $O^-(g)$	+844

$\Delta H^{\ominus}_{\text{latt}}$ $Na_2O(s)$ = ... kJ mol^{-1} [4]

(e) State how $\Delta H^{\ominus}_{\text{latt}}$ $Na_2S(s)$ differs from $\Delta H^{\ominus}_{\text{latt}}$ $Na_2O(s)$.
Indicate this by placing a tick (✓) in the appropriate box in the table.

$\Delta H^{\ominus}_{\text{latt}}$ $Na_2S(s)$ is more exothermic than $\Delta H^{\ominus}_{\text{latt}}$ $Na_2O(s)$	$\Delta H^{\ominus}_{\text{latt}}$ $Na_2S(s)$ is the same as $\Delta H^{\ominus}_{\text{latt}}$ $Na_2O(s)$	$\Delta H^{\ominus}_{\text{latt}}$ $Na_2S(s)$ is less exothermic than $\Delta H^{\ominus}_{\text{latt}}$ $Na_2O(s)$

Explain your answer.

....

....

....
[2]

[Total: 15]

19. 9701/41/M/J/18 Q1c

(c) **(i)** Silicon tetrachloride can be prepared according to reaction 1.

reaction 1 $Si(s) + 2Cl_2(g) \rightarrow SiCl_4(l)$ $\Delta S^\ominus = -225.7\,J\,K^{-1}\,mol^{-1}$

standard entropy of silicon, $S^\ominus$ Si(s)	18.7 J K^{-1} mol^{-1}
standard entropy of silicon tetrachloride, $S^\ominus$ $SiCl_4$(l)	239.0 J K^{-1} mol^{-1}

Calculate the standard entropy of chlorine, $S^\ominus$ Cl_2(g). Show all your working.

$S^\ominus$ Cl_2(g) = ... J K^{-1} mol^{-1} [2]

(ii) Explain why the entropy change for reaction 1 is negative.

....

.... [1]

(d) The standard enthalpy change of formation of silicon tetrachloride, $\Delta H_f^\ominus$ $SiCl_4$(l), is –640 kJ mol^{-1}.

Reaction 1 is spontaneous at lower temperatures, but it is not spontaneous at very high temperatures.

Calculate the temperature above which reaction 1 is **not** spontaneous.

temperature = ... K [2]

[Total: 13]

20. 9701/41/O/N/19 Q3

(a) Explain what is meant by the term *entropy of a system*.

....

.... [1]

(b) State and explain whether the entropy change of each of the following processes is positive or negative. Do not consider the entropy change of the surroundings.

- liquid water at 80 °C is cooled to 60 °C

The entropy change is .... because ....

....

- solid calcium chloride is added to water and the mixture is stirred

The entropy change is .... because ....

....

- the change corresponding to the lattice energy of calcium chloride, ΔH_{latt} $CaCl_2(s)$, takes place

The entropy change is .... because ....

....

[3]

(c) The reaction $ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$ is not spontaneous at room temperature.

(i) Give the full name for the term ΔG°.

.... [1]

(ii) Describe how the temperature at which the reaction becomes spontaneous can be calculated. Include an equation in your answer.

equation ....

....

....

....

[2]

[Total: 7]

21. 9701/42/O/N/19 Q4

The table shows some standard entropy data.

substance	standard entropy, $S^\ominus$ / $\text{J K}^{-1}\text{mol}^{-1}$
PbO_2(s)	77
PbO(s)	69
O_2(g)	205

Lead(IV) oxide, PbO_2, decomposes to lead(II) oxide, PbO, and oxygen when heated.

$$2PbO_2(s) \rightarrow 2PbO(s) + O_2(g) \qquad \Delta H^\ominus = +118\,\text{kJ mol}^{-1}$$

(a) Use the data to calculate the value of $\Delta S^\ominus$ for this reaction.

$\Delta S^\ominus$ = ... $\text{J K}^{-1}\text{mol}^{-1}$ [2]

(b) Use the value of $\Delta H^\ominus$ and your answer to **(a)** to calculate the temperature at which this reaction becomes feasible.

T = ... K [3]

(c) Solid lead(II) oxide can be made by heating lead metal in air.

Predict the **sign** of the standard entropy change of this reaction. Explain your answer.

....

.... [1]

[Total: 6]

22. 9701/41/M/J/19 Q6

(a) Complete the table by placing **one** tick (✓) in each row to indicate the sign of each type of energy change under standard conditions.

energy change	always positive	always negative	either negative or positive
bond energy			
enthalpy change of formation			

[1]

(b) Explain what is meant by the term *enthalpy change of atomisation*.

....

.... [1]

(c) The overall reaction for the atomisation of liquid bromine molecules, $Br_2(l)$, is shown.

$$Br_2(l) \rightarrow 2Br(g)$$

This happens via a two-step process.

- Construct a **labelled** energy cycle to represent this atomisation process, including state symbols.
- Use your cycle and relevant data from the *Data Booklet* to calculate the enthalpy change of vaporisation of $Br_2(l)$, $\Delta H^{\ominus}_{vap}$.
  The enthalpy change of atomisation of bromine, ΔH_{at}, = +112 kJ mol^{-1}.

$\Delta H^{\ominus}_{vap}$ = ... kJ mol^{-1} [3]

(d) Suggest how the $\Delta H^{\ominus}_{vap}$ of iodine, $I_2(l)$, would compare to that of bromine, $Br_2(l)$. Explain your answer.

....

.... [1]

(e) **(i)** Explain what is meant by the term *enthalpy change of hydration*.

....

.... [1]

(ii) Suggest why the enthalpy change of hydration of $Br^-(g)$ is **more** exothermic than that of $I^-(g)$.

....

....

.... [2]

[Total: 9]

23. 9701/42/M/J/19 Q5

(a) Complete the table by placing **one** tick (✓) in each row to indicate the sign of each type of energy change under standard conditions.

energy change	always positive	always negative	either negative or positive
lattice energy			
enthalpy change of neutralisation			

[1]

(b) Define, in words, the term *enthalpy change of solution*.

....

.... [1]

(d) Some lattice energy values are shown in the table.

compound	lattice energy value / kJ mol^{-1}
$CaBr_2$(s)	–2176
KBr(s)	–679

Suggest an explanation for why ΔH°_{latt} $CaBr_2$ is **more** exothermic than ΔH°_{latt} KBr.

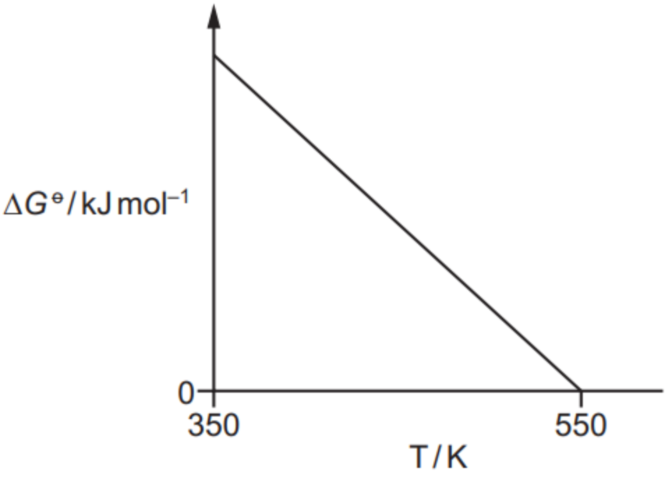
....

....

.... [2]

(e) For a particular gas phase reaction the variation in standard Gibbs free energy change, ΔG°, with temperature is shown.

Assume standard enthalpy change, ΔH°, and standard entropy change, ΔS°, remain constant with temperature.



(i) Write the equation that relates ΔG° to ΔH° and ΔS°.

.... [1]

(ii) Use this equation to explain why ΔG° becomes **less** positive as temperature increases in this reaction.

....

.... [1]

[Total: 9]

24. 9701/43/M/J/19 Q6

(a) Complete the table by placing **one** tick (✓) in each row to indicate the sign of each type of energy change under standard conditions.

energy change	always positive	always negative	either negative or positive
bond energy			
enthalpy change of formation			

[1]

(b) Explain what is meant by the term *enthalpy change of atomisation*.

....

.... [1]

(c) The overall reaction for the atomisation of liquid bromine molecules, $Br_2(l)$, is shown.

$$Br_2(l) \rightarrow 2Br(g)$$

This happens via a two-step process.

- Construct a **labelled** energy cycle to represent this atomisation process, including state symbols.
- Use your cycle and relevant data from the *Data Booklet* to calculate the enthalpy change of vaporisation of $Br_2(l)$, $\Delta H^{\ominus}_{vap}$.
  The enthalpy change of atomisation of bromine, ΔH_{at}, = +112 kJ mol^{-1}.

$\Delta H^{\ominus}_{vap}$ = ... kJ mol^{-1} [3]

(d) Suggest how the $\Delta H^{\ominus}_{vap}$ of iodine, $I_2(l)$, would compare to that of bromine, $Br_2(l)$. Explain your answer.

....

.... [1]

(e) **(i)** Explain what is meant by the term *enthalpy change of hydration*.

....

.... [1]

(ii) Suggest why the enthalpy change of hydration of $Br^-(g)$ is **more** exothermic than that of $I^-(g)$.

....

....

.... [2]

[Total: 9]

25. 9701/42/F/M/19 Q2 (c)

(c) The equation for the formation of the gaseous hydroxide ion is shown.

$$\tfrac{1}{2}H_2(g) + \tfrac{1}{2}O_2(g) + e^- \rightarrow OH^-(g) \qquad \Delta H = \Delta H_f^\ominus(OH^-(g))$$

Use data in the table and from the *Data Booklet* to calculate $\Delta H_f^\ominus(OH^-(g))$. You might find it useful to construct a Born-Haber cycle.

enthalpy change	$\Delta H^\ominus / kJ\,mol^{-1}$
atomisation of Mg(s)	+148
formation of $Mg(OH)_2(s)$	–925
lattice energy of $Mg(OH)_2(s)$	–2993

$\Delta H_f^\ominus(OH^-(g))$ = ... $kJ\,mol^{-1}$ [3]

26. 9701/41/O/N/20 2c

(c) Use the following data and relevant data from the *Data Booklet* to calculate a value for the lattice energy of magnesium fluoride, $MgF_2(s)$.

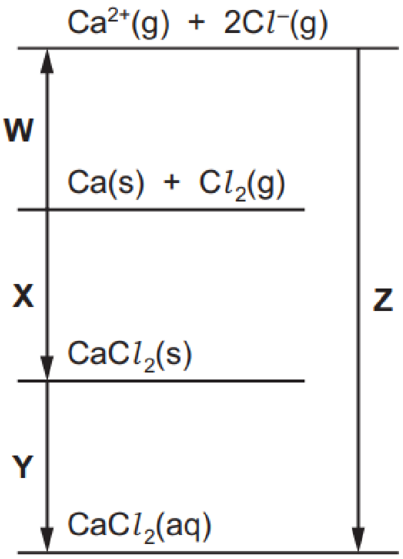
You might find it helpful to construct an energy cycle.
Show your working.

electron affinity of F(g) = –348 kJ mol^{-1}
enthalpy change of atomisation of Mg(s) = +147 kJ mol^{-1}
enthalpy change of formation of $MgF_2(s)$ = –1102 kJ mol^{-1}

$\Delta H_{latt} MgF_2(s)$ = ... [3]

27. 9701/42/O/N/20 3 (a)

(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]

28. 9701/42/M/J/20 Q7

(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]

(c) The lattice energies of $ZnBr_2$, $ZnCl_2$ and ZnO are shown.

compound	lattice energy / kJ mol^{-1}
$ZnBr_2$	–2678
$ZnCl_2$	–2734
ZnO	–3971

(i) Explain why there is a difference between the lattice energies of $ZnBr_2$ and $ZnCl_2$.

....

.... [1]

(ii) Explain why there is a difference between the lattice energies of $ZnCl_2$ and ZnO.

....

.... [1]

[Total: 10]

29. 9701/43/M/J/20 Q7(b)

(b) Silver chloride, AgC*l*, is sparingly soluble in water. The equation for the enthalpy change of solution is shown.

$$AgCl(s) \rightarrow Ag^{+}(aq) + Cl^{-}(aq) \qquad \Delta H^{\ominus}_{sol} = +65.5\,kJ\,mol^{-1}$$

Standard entropies are shown in the table.

species	AgC*l*(s)	$Ag^{+}(aq)$	$Cl^{-}(aq)$
$S^{\ominus}$ / J K^{-1} mol^{-1}	+96.2	+72.7	+56.5

(i) Calculate the standard entropy change of solution, $\Delta S^{\ominus}$.

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

(ii) Explain, with the aid of a calculation, why AgC*l* is insoluble in water at 25 °C.

You should use data from this question and your answer to **(b)(i)**.

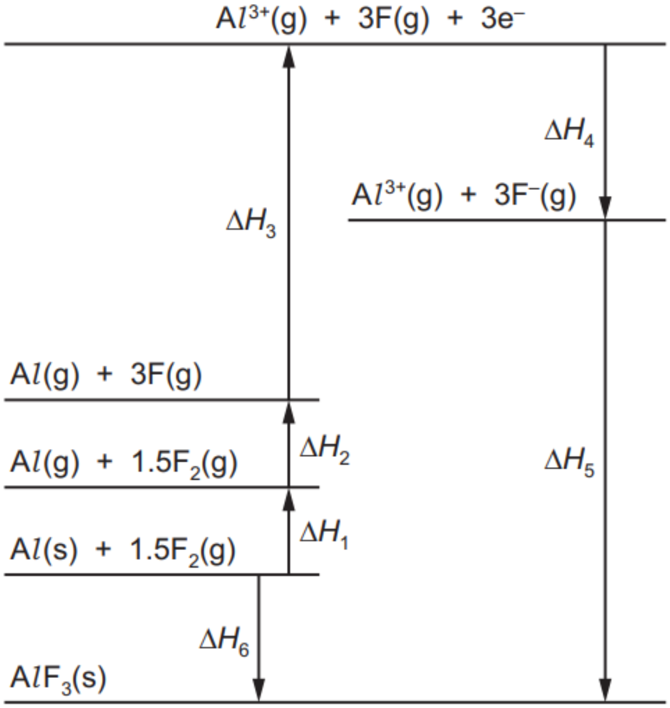
....

.... [3]

30. 9701/42/F/M/20 Q3(c)

(c) AlF_3 is an ionic compound.

The Born–Haber cycle for the formation of AlF_3 is shown.



(i) Name the enthalpy changes labelled ΔH_4 and ΔH_6.

ΔH_4 = ...

ΔH_6 = ...

[2]

(ii) Use the data in the table and data from the *Data Booklet* to calculate the lattice energy of AlF_3.

process	enthalpy change / kJ mol^{-1}
$Al(s) \rightarrow Al(g)$	+326
$Al(g) \rightarrow Al^{3+}(g)$	+5137
$F(g) \rightarrow F^-(g)$	–328
$Al(s) + 1.5F_2(g) \rightarrow AlF_3(s)$	–1504

lattice energy of AlF_3 = ... kJ mol^{-1} [2]

(iii) Scandium fluoride, ScF_3, is an ionic compound.

Use data from the *Data Booklet* to suggest how the lattice energy of AlF_3 compares with the lattice energy of ScF_3.

Explain your answer.

....

....

.... [2]

31. 9701/41/O/N/21 Q3

(a) Define the term *electron affinity*.

....

.... [2]

(b) Write an equation for the process corresponding to the **second** ionisation energy of calcium. Include state symbols.

.... [1]

Some data relating to calcium and oxygen are listed. Select relevant data from this list for your answers to parts **(c)**, **(d)** and **(e)**.

process	value / kJ mol^{-1}
first ionisation energy of oxygen	+1310
second ionisation energy of oxygen	+3390
first electron affinity of oxygen	–142
second electron affinity of oxygen	+844
enthalpy change for $\frac{1}{2}O_2(g) + 2e^- \rightarrow O^{2-}(g)$	+951
enthalpy change for $Ca(s) \rightarrow Ca^{2+}(g) + 2e^-$	+1933
lattice energy of CaO(s)	–3517

(c) Oxygen exists as O_2 molecules.

Use the data in this question to calculate a value for the bond energy of the O=O bond.
Show all your working.

bond energy = ... kJ mol^{-1} [3]

(d) **(i)** Suggest why the first electron affinity of oxygen is negative.

....

.... [1]

(ii) Suggest why the second electron affinity of oxygen is positive.

....

.... [1]

(e) Calculate the enthalpy of formation of calcium oxide, CaO(s).

enthalpy of formation = ... kJ mol^{-1} [2]

(f) The lattice energy of lithium fluoride, LiF(s), is –1022 kJ mol^{-1}.

Identify the factor that causes the lattice energy of calcium oxide to be more exothermic than that of lithium fluoride. Explain why this factor causes the difference in lattice energies.

....

.... [2]

[Total: 12]

32. 9701/42/O/N/21 Q1

Radium is a Group 2 element.

The predicted lattice energy, $\Delta H^{\ominus}_{\text{latt}}$, of radium sulfide, RaS, is –2612 kJ mol⁻¹.

(a) Define $\Delta H^{\ominus}_{\text{latt}}$.

....

.... [2]

Some data relating to radium and sulfur are listed. Select relevant data from this list for use in your answers to parts **(b)** to **(e)**.

process	value / kJ mol^{-1}
enthalpy change for $Ra(s) \rightarrow Ra^{2+}(g) + 2e^-$	+1619
first ionisation energy of sulfur	+1000
second ionisation energy of sulfur	+2260
first electron affinity of sulfur	–200
second electron affinity of sulfur	+532
enthalpy change for $\frac{1}{8}S_8(s) + 2e^- \rightarrow S^{2-}(g)$	+555
lattice energy of RaS(s)	–2612

(b) Write an equation for the process corresponding to the **second** electron affinity of sulfur. Include state symbols.

.... [1]

(c) Sulfur exists as S_8 molecules in the solid state.

Use the data in this question to calculate the enthalpy change for the reaction $S_8(s) \rightarrow 8S(g)$.

enthalpy change = ... kJ mol^{-1} [3]

(d) Calculate the standard enthalpy change of formation, $\Delta H_f^\ominus$, of radium sulfide.

standard enthalpy change, $\Delta H_f^\ominus$ = ... kJ mol^{-1} [2]

(e) **(i)** State the **two** major factors that affect the numerical magnitude of a lattice energy.

....

.... [2]

(ii) For **each** factor you have identified in **(e)(i)**, state whether it tends to make the lattice energy of radium sulfide more or less exothermic than that of sodium chloride.

Explain your answer.

....

....

....

.... [2]

(iii) The lattice energies of sodium chloride, NaC*l*, and radium sulfide, RaS, are –771 kJ mol^{-1} and –2612 kJ mol^{-1}, respectively.

Identify the **dominant** factor in determining the relative numerical magnitudes of the lattice energies of radium sulfide and sodium chloride.

Explain your answer.

....

.... [1]

[Total: 13]

33. 9701/42/O/N/21 3 (e)

(e) The standard entropies, $S^\ominus$, of three species are given in the table.

species	$S^\ominus$ / J K^{-1} mol^{-1}
$HCl(g)$	+187
$H_2(g)$	+131
$Cl_2(g)$	+223

(i) Calculate $\Delta S^\ominus$ for the reaction $2HCl(g) \rightarrow H_2(g) + Cl_2(g)$.

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

(ii) $\Delta H^\ominus$ for the reaction $2HCl(g) \rightarrow H_2(g) + Cl_2(g)$ is +185 kJ mol^{-1}.

Calculate $\Delta G^\ominus$ for this reaction at 298 K.

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

(iii) Predict the effect of increasing temperature on the spontaneity of this reaction. Explain your answer.

....

.... [1]

34. 9701/42/M/J/21 Q4 (a), (c), (d)

(a) **(i)** Define the term *lattice energy*.

....

....

.... [2]

(ii) Use the following data to calculate a value for the enthalpy change of solution of copper(II) chloride, $CuCl_2(s)$. You might find it helpful to construct an energy cycle.

enthalpy change of hydration of Cl^- = $-378\,kJ\,mol^{-1}$
enthalpy change of hydration of Cu^{2+} = $-2099\,kJ\,mol^{-1}$
lattice energy of $CuCl_2(s)$ = $-2824\,kJ\,mol^{-1}$

enthalpy change of solution of $CuCl_2(s)$ = ... $kJ\,mol^{-1}$ [2]

(iii) The enthalpy change of hydration of Ca^{2+} is $-1579\,kJ\,mol^{-1}$.

Use the *Data Booklet* to suggest why there is a big difference in the values of ΔH_{hyd} for Ca^{2+} and Cu^{2+}.

....

....

.... [2]

(c) **(i)** Explain what is meant by the term *entropy of a system*.

....

.... [1]

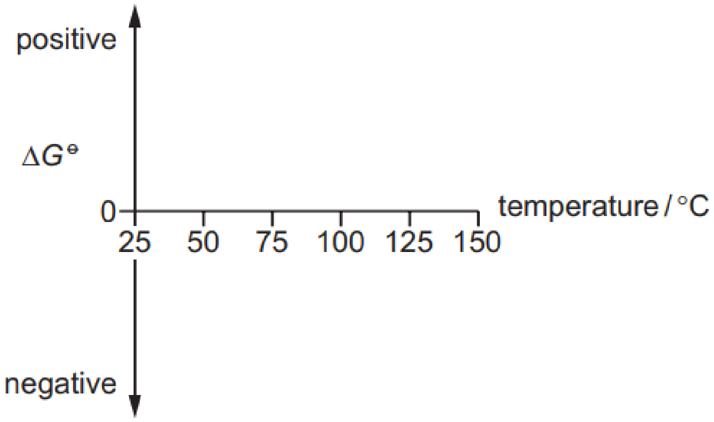
(ii) Place one tick (✓) in each row of the table to show the sign of each entropy change, Δ*S*.

process	Δ*S* is negative	Δ*S* is zero	Δ*S* is positive
NaC*l* dissolving in water			
water solidifying to ice			

[1]

(iii) The evaporation of one mole of water has a standard Gibbs free energy change, $\Delta G^{\ominus}$, of +8.6 kJ at 25 °C.

Sketch a graph on the axes to show how $\Delta G^{\ominus}$ changes for this process between 25 °C and 150 °C at 101 kPa.



[2]

(d) The reaction between A and B is feasible at low temperatures but is **not** feasible at high temperatures.

$$A + B \rightleftharpoons C + D$$

Deduce the signs of ΔH and ΔS for this reaction and explain why the feasibility changes with temperature.

sign of ΔH = ... sign of ΔS = ...

....

....

....

.... [2]

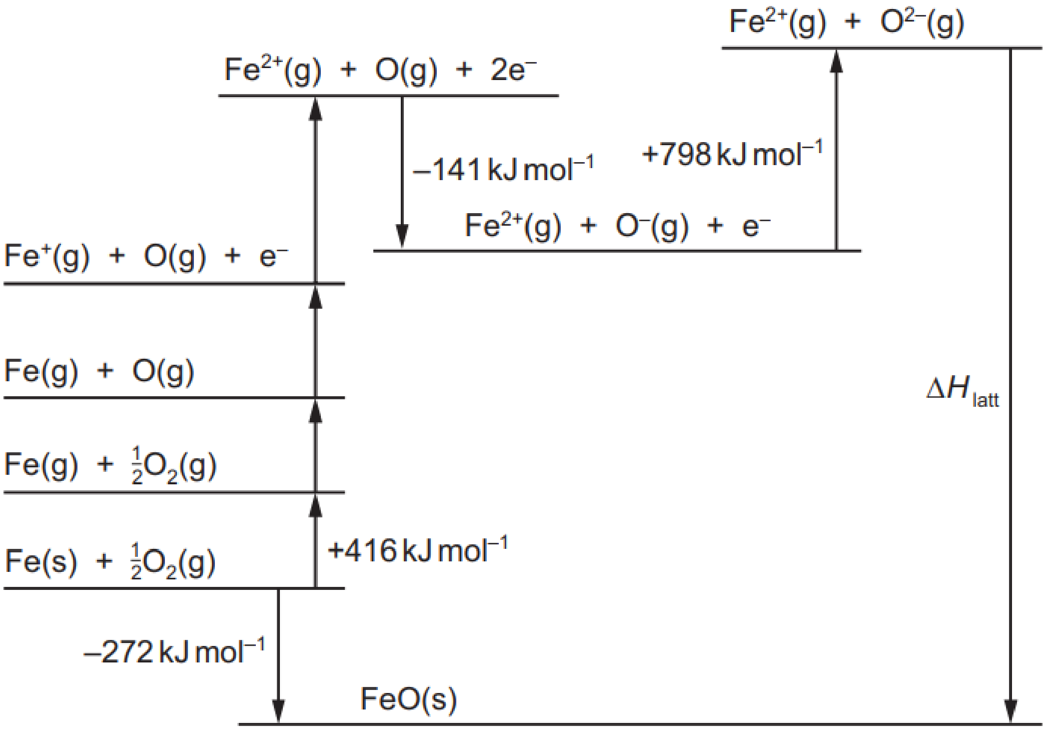
[Total: 15]

35. 9701/42/F/M/21 Q2 a(i) (iii)

(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]

(iii) State and explain how the lattice energy of FeO(s) compares to the lattice energy of CaO(s).

....

....

....

....

.... [2]

36. 9701/41/O/N/22 Q1

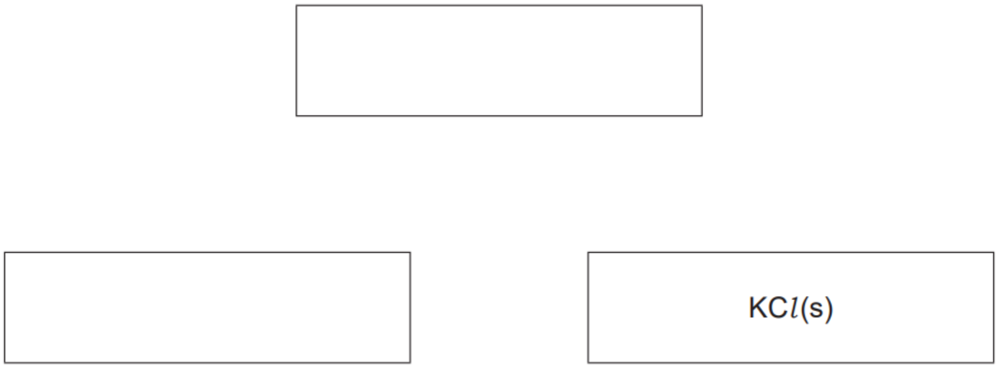
Potassium chloride, $\mathrm{KC}l$, and magnesium chloride, $\mathrm{MgC}l_2$, are both ionic solids.

Table 1.1

energy change	value / kJ mol^{-1}
standard enthalpy change of solution, $\Delta H^{\ominus}_{\mathrm{sol}}$, of $\mathrm{KC}l$	+15
lattice energy, $\Delta H^{\ominus}_{\mathrm{latt}}$, of $\mathrm{KC}l(\mathrm{s})$	–701
standard enthalpy change of hydration, $\Delta H^{\ominus}_{\mathrm{hyd}}$, of K^+	–322
standard enthalpy change of hydration, $\Delta H^{\ominus}_{\mathrm{hyd}}$, of $\mathrm{C}l^-$	–364
standard enthalpy change of solution, $\Delta H^{\ominus}_{\mathrm{sol}}$, of $\mathrm{MgC}l_2$	–155
lattice energy, $\Delta H^{\ominus}_{\mathrm{latt}}$, of $\mathrm{MgC}l_2(\mathrm{s})$	–2493

(a) Complete the energy cycle involving the enthalpy change of solution and the lattice energy of potassium chloride, $\mathrm{KC}l$, and the relevant enthalpy changes of hydration. Label your diagram.

State symbols should be used.



[2]

(b) Use the data in Table 1.1 to calculate the enthalpy change of hydration of magnesium ions, Mg^{2+}. Show your working.

$\Delta H^{\ominus}_{\mathrm{hyd}}$ of magnesium ions, Mg^{2+} = ... kJ mol^{-1} [2]

(c) Explain the reasons why the lattice energy of $MgCl_2$ is more exothermic than the lattice energy of KCl.

....

....

....

.... [2]

(d) Define the following terms.

(i) enthalpy change of atomisation ....

.... [1]

(ii) first electron affinity ....

.... [1]

(e) **(i)** Explain what is meant by entropy, *S*.

....

.... [1]

(ii) Potassium chloride is very soluble in water at 20 °C.

Explain the solubility of potassium chloride by reference to change in entropy, ΔS.

....

.... [1]

(iii) Use the Gibbs equation and your answer to **(e)(ii)** to predict whether potassium chloride is more soluble in water at 20 °C or at 80 °C. Explain your answer.

....

....

.... [1]

[Total: 11]

37. 9701/42/O/N/22 Q1

Calcium chloride, $CaCl_2$, is an ionic solid.

The values of some energy changes are shown in Table 1.1.

Table 1.1

energy change	value / kJ mol^{-1}
lattice energy, $\Delta H^{\ominus}_{\text{latt}}$, $CaCl_2(s)$	–2237
standard enthalpy change of atomisation of calcium	+193
first ionisation energy of calcium	+590
second ionisation energy of calcium	+1150
standard enthalpy change of atomisation of chlorine	+121
first electron affinity of chlorine	–364

(a) Define lattice energy.

....

.... [1]

(b) Use the data in Table 1.1 to calculate the standard enthalpy change of formation, $\Delta H^{\ominus}_{f}$, of calcium chloride. It may be helpful to draw an energy cycle. Show all your working.

$\Delta H^{\ominus}_{f}(CaCl_2(s))$ = ... kJ mol^{-1} [2]

(c) Three possible values for the first electron affinity of bromine are shown in Table 1.2. One of them is correct.

Place a tick by the correct value. Explain your choice.

Table 1.2

possible values	place **one** tick (✓) in this column
$-342\,\text{kJ mol}^{-1}$	
$-364\,\text{kJ mol}^{-1}$	
$-386\,\text{kJ mol}^{-1}$	

explanation ....

.... [1]

(d) The enthalpy change of hydration of the chloride ion can be calculated using the lattice energy of calcium chloride and the data shown in Table 1.3.

Table 1.3

energy change	value / kJ mol^{-1}
standard enthalpy change of solution of $CaCl_2$(s)	–83
standard enthalpy change of hydration of Ca^{2+}(g)	–1650

(i) Define the following terms.

enthalpy change of solution ....

....

enthalpy change of hydration ....

....
[2]

(ii) Calculate the standard enthalpy change of hydration of the chloride ion, Cl^-(g). It may be helpful to draw an energy cycle. Show all your working.

$\Delta H^{\ominus}_{hyd}(Cl^-(g))$ = ... kJ mol^{-1} [2]

(e) Calcium fluoride, $CaF_2(s)$, can be synthesised directly from its elements.

The value of $\Delta H_f^\ominus(CaF_2(s))$ is –1214 kJ mol^{-1}.

(i) Predict the sign of the entropy change, $\Delta S^\ominus$, for this synthesis. Explain your answer.

The sign of the entropy change is ....

explanation ....

....

[1]

(ii) Use the value of $\Delta H_f^\ominus(CaF_2(s))$ given in **(e)** and your answer to **(e)(i)** to predict how the feasibility for this synthesis will change with increasing temperature.

....

....

....

.... [2]

[Total: 11]

38. 9701/42/M/J/22 Q3

(a) **(i)** Define first electron affinity.

....

....

.... [2]

(ii) The first electron affinity of an atom is usually an exothermic process, whereas the second electron affinity is an endothermic process.

Suggest why.

....

.... [1]

(iii) Describe the general trend in first electron affinities for C*l*, Br and I. Explain your answer.

....

....

.... [2]

(b) Table 3.1 shows energy changes to be used in this question and in **(c)**.

Table 3.1

energy change	value / kJ mol^{-1}
standard enthalpy change of atomisation of zinc	+131
first ionisation energy of zinc	+906
second ionisation energy of zinc	+1733
standard enthalpy change of formation of $ZnI_2(s)$	–208
lattice energy, $\Delta H^{\ominus}_{latt}$, of zinc iodide, $ZnI_2(s)$	–2605
first ionisation energy of iodine	+1008
second ionisation energy of iodine	+1846
I–I bond energy	+151
enthalpy change of sublimation of iodine, $I_2(s) \rightarrow I_2(g)$	+62

Calculate the first electron affinity for iodine. Use relevant data from Table 3.1 in your working.
It may be helpful to draw a labelled energy cycle.
Show all working.

first electron affinity for iodine = ... kJ mol^{-1} [3]

(c) Predict how $\Delta H^{\ominus}_{latt}$ of $CdI_2(s)$ differs from $\Delta H^{\ominus}_{latt}$ of $ZnI_2(s)$.
Place a tick (✓) in the appropriate box in Table 3.2.

Table 3.2

$\Delta H^{\ominus}_{latt}$ of $CdI_2(s)$ is less negative than $\Delta H^{\ominus}_{latt}$ of $ZnI_2(s)$	$\Delta H^{\ominus}_{latt}$ of $CdI_2(s)$ is the same as $\Delta H^{\ominus}_{latt}$ of $ZnI_2(s)$	$\Delta H^{\ominus}_{latt}$ of $CdI_2(s)$ is more negative than $\Delta H^{\ominus}_{latt}$ of $ZnI_2(s)$

Explain your answer.

....

.... [1]

[Total: 9]

39. 9701/42/M/J/22 Q4(a)

(a) Calcium carbonate decomposes on heating.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

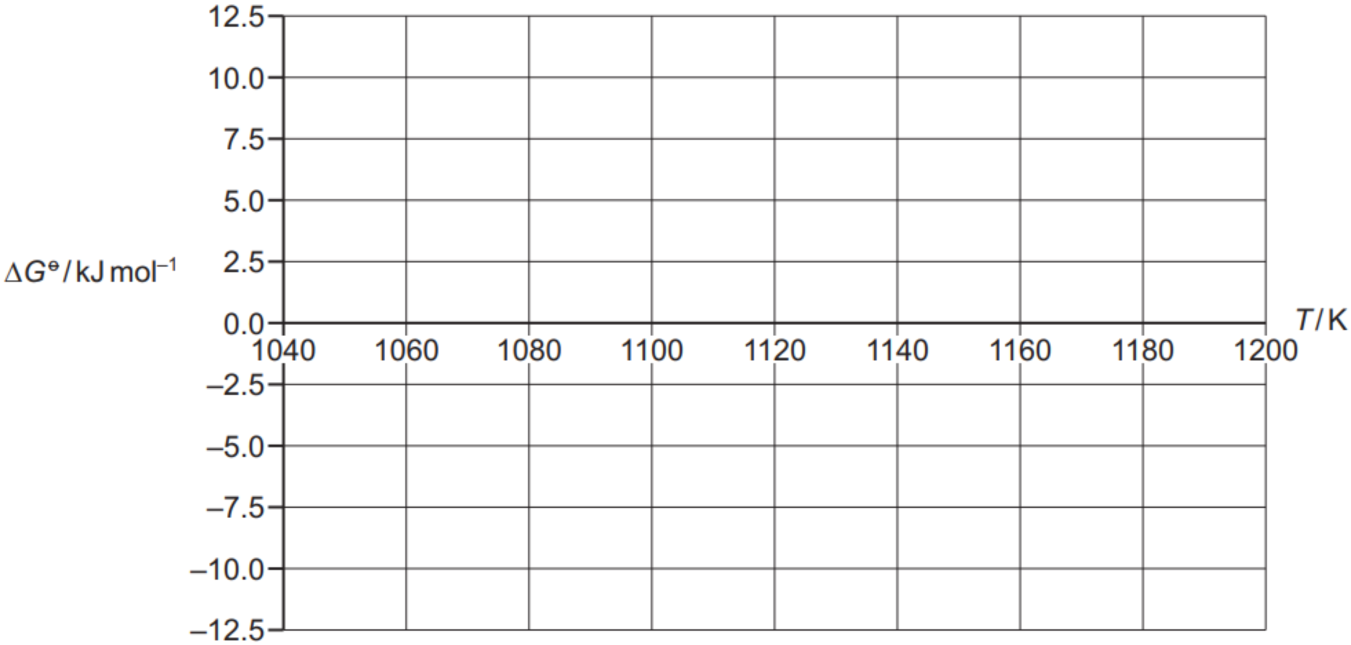
Table 4.1 shows the values of the Gibbs free energy change, $\Delta G^{\ominus}$, for this reaction at various temperatures.

Table 4.1

T/K	$\Delta G^{\ominus}$/kJ mol^{-1}
1050	9.9
1085	4.3
1120	–1.3
1148	–5.8
1176	–10.3

Assume the standard enthalpy change, $\Delta H^{\ominus}$, and the standard entropy change, $\Delta S^{\ominus}$, for this reaction remain constant over this temperature range.

(i) Use the data in Table 4.1 to plot a graph of $\Delta G^{\ominus}$ against T on the grid.



[2]

(ii) Calculate the gradient of your graph. Determine the $\Delta S^{\ominus}$ in J K^{-1} mol^{-1} for this reaction. Show all working.

$\Delta S^{\ominus}$ = ... J K^{-1} mol^{-1} [2]

40. 9701/43/M/J/22 Q1 (d)

(d) **(i)** The equation for the formation of a gaseous sulfate ion is shown.

$$S(s) + 2O_2(g) + 2e^- \rightarrow SO_4^{2-}(g) \qquad \Delta H = \Delta H_f^{\ominus} \text{ of } SO_4^{2-}(g)$$

Calculate the standard enthalpy change of formation, $\Delta H_f^{\ominus}$, of $SO_4^{2-}(g)$. It may be helpful to draw a labelled energy cycle. Use relevant data from Table 1.1 in your calculations.

Table 1.1

energy change	value / kJ mol^{-1}
lattice energy of barium sulfate, $BaSO_4(s)$	–2469
standard enthalpy change of formation of barium sulfate	–1473
standard enthalpy change of atomisation of barium	+180
first ionisation energy of barium	+503
second ionisation energy of barium	+965
standard enthalpy change of atomisation of sulfur	+279
standard enthalpy change for $S(g) \rightarrow S^{2-}(g)$	+440
standard enthalpy change for $O(g) \rightarrow O^{2-}(g)$	+657
O=O bond energy	+496

$\Delta H_f^{\ominus}$ of $SO_4^{2-}(g)$ = ... kJ mol^{-1} [3]

(ii) Suggest how the lattice energy of $BaSO_4(s)$ differs from the lattice energy of $Cs_2SO_4(s)$. Explain your answer.

....

....

.... [2]

(e) The reaction of solid hydrated barium hydroxide, $Ba(OH)_2 \cdot 8H_2O$, with ammonium salts is endothermic.

(i) Calculate the **minimum** temperature at which the reaction of $Ba(OH)_2 \cdot 8H_2O$ with NH_4NO_3 becomes feasible. Show all your working.

$$Ba(OH)_2 \cdot 8H_2O(s) + 2NH_4NO_3(s) \rightarrow 2NH_3(g) + Ba(NO_3)_2(s) + 10H_2O(l) \quad \Delta H_r^\ominus = +132\,\text{kJ mol}^{-1}$$

$$\Delta S^\ominus = +616\,\text{J K}^{-1}\,\text{mol}^{-1}$$

temperature = ... °C [2]

(ii) Barium hydroxide reacts readily with ammonium chloride on mixing at room temperature.

$$Ba(OH)_2 \cdot 8H_2O(s) + 2NH_4Cl(s) \rightarrow 2NH_3(g) + BaCl_2 \cdot 2H_2O(s) + 8H_2O(l) \quad \Delta H_r^\ominus = +133\,\text{kJ mol}^{-1}$$

Some relevant standard entropies are given in Table 1.2.

Table 1.2

substance	$Ba(OH)_2 \cdot 8H_2O(s)$	$NH_4Cl(s)$	$NH_3(g)$	$BaCl_2 \cdot 2H_2O(s)$	$H_2O(l)$
$S^\ominus$ / J K^{-1} mol^{-1}	427	95	192	203	70

Calculate the standard Gibbs free energy change, $\Delta G^\ominus$, for this reaction at 25 °C.

$\Delta G^\ominus$ = ... kJ mol^{-1} [3]

[Total: 16]

41. 9701/41/M/J/22 Q 1(d)

(d) (i) The equation for the formation of a gaseous sulfate ion is shown.

$$S(s) + 2O_2(g) + 2e^- \rightarrow SO_4^{2-}(g) \qquad \Delta H = \Delta H_f^{\ominus} \text{ of } SO_4^{2-}(g)$$

Calculate the standard enthalpy change of formation, $\Delta H_f^{\ominus}$, of $SO_4^{2-}(g)$. It may be helpful to draw a labelled energy cycle. Use relevant data from Table 1.1 in your calculations.

Table 1.1

energy change	value / kJ mol^{-1}
lattice energy of barium sulfate, $BaSO_4(s)$	–2469
standard enthalpy change of formation of barium sulfate	–1473
standard enthalpy change of atomisation of barium	+180
first ionisation energy of barium	+503
second ionisation energy of barium	+965
standard enthalpy change of atomisation of sulfur	+279
standard enthalpy change for $S(g) \rightarrow S^{2-}(g)$	+440
standard enthalpy change for $O(g) \rightarrow O^{2-}(g)$	+657
O=O bond energy	+496

$\Delta H_f^{\ominus}$ of $SO_4^{2-}(g)$ = ... kJ mol^{-1} [3]

(ii) Suggest how the lattice energy of $BaSO_4(s)$ differs from the lattice energy of $Cs_2SO_4(s)$. Explain your answer.

....

....

.... [2]

(e) The reaction of solid hydrated barium hydroxide, $Ba(OH)_2 \cdot 8H_2O$, with ammonium salts is endothermic.

(i) Calculate the **minimum** temperature at which the reaction of $Ba(OH)_2 \cdot 8H_2O$ with NH_4NO_3 becomes feasible. Show all your working.

$$Ba(OH)_2 \cdot 8H_2O(s) + 2NH_4NO_3(s) \rightarrow 2NH_3(g) + Ba(NO_3)_2(s) + 10H_2O(l) \quad \Delta H_r^\ominus = +132\,kJ\,mol^{-1}$$

$$\Delta S^\ominus = +616\,J\,K^{-1}\,mol^{-1}$$

temperature = ... °C [2]

(ii) Barium hydroxide reacts readily with ammonium chloride on mixing at room temperature.

$$Ba(OH)_2 \cdot 8H_2O(s) + 2NH_4Cl(s) \rightarrow 2NH_3(g) + BaCl_2 \cdot 2H_2O(s) + 8H_2O(l) \quad \Delta H_r^\ominus = +133\,kJ\,mol^{-1}$$

Some relevant standard entropies are given in Table 1.2.

Table 1.2

substance	$Ba(OH)_2 \cdot 8H_2O(s)$	$NH_4Cl(s)$	$NH_3(g)$	$BaCl_2 \cdot 2H_2O(s)$	$H_2O(l)$
$S^\ominus / J\,K^{-1}\,mol^{-1}$	427	95	192	203	70

Calculate the standard Gibbs free energy change, $\Delta G^\ominus$, for this reaction at 25 °C.

$\Delta G^\ominus$ = ... $kJ\,mol^{-1}$ [3]

[Total: 16]

42. 9701/42/F/M/22 Q 2 a,b

Silicon is the second most abundant element by mass in the Earth's crust.

(a) In industry, silicon is extracted from SiO_2 by reaction with carbon at over 2000 °C.

reaction 1 $SiO_2(s) + 2C(s) \rightarrow Si(l) + 2CO(g)$

(i) Explain why the entropy change, ΔS, of reaction 1 is positive.

....

.... [1]

(ii) Reaction 1 is highly endothermic.

Suggest the effect of an increase in temperature on the feasibility of this reaction.
Explain your answer.

....

....

.... [2]

(b) Silicon is purified by first heating it in a stream of $HCl(g)$ to form $SiHCl_3$. The $SiHCl_3$ formed is then distilled to remove other impurities.

reaction 2 $Si(s) + 3HCl(g) \rightarrow SiHCl_3(g) + H_2(g)$

(i) Table 2.1 shows some standard entropy data.

Table 2.1

compound	standard entropy, $S^\ominus / J K^{-1} mol^{-1}$
$Si(s)$	19
$HCl(g)$	187
$SiHCl_3(g)$	314
$H_2(g)$	131

Use the data in Table 2.1 to calculate $\Delta S^\ominus$ for reaction 2.

$\Delta S^\ominus$ = ... $J K^{-1} mol^{-1}$ [2]

(ii) Reaction 3 is the reverse of reaction 2 and is used to obtain pure silicon.

reaction 3 $SiHCl_3(g) + H_2(g) \rightarrow Si(s) + 3HCl(g)$ $\Delta H = +219.3\,kJ\,mol^{-1}$

Use this information and your answer to **(b)(i)** to calculate the temperature, in K, at which reaction 3 becomes feasible.
Show your working.

[If you were unable to answer **(b)(i)**, you should use $\Delta S^{\ominus} = -150\,J\,K^{-1}\,mol^{-1}$ for reaction 2. This is not the correct answer to **(b)(i)**.]

temperature = ... K [2]