

1(a) Two changes are described below.

For each change,

- write an equation, including state symbols,
- state and explain how the entropy changes.

(i) The reaction of aqueous barium nitrate with aqueous sodium sulfate.

Full equation with state symbols

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Explanation of entropy change

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.....

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

(ii) The change that accompanies the standard enthalpy change of atomisation of iodine.

Equation with state symbols

.....

Explanation of entropy change

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.....

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

2 This question is about free energy changes, ΔG , enthalpy changes, ΔH , and temperature, T .

(a) The Gibbs' equation is shown below.

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

A chemist investigates a reaction to determine how ΔG varies with T . The results are shown in **Fig. 18.1**.

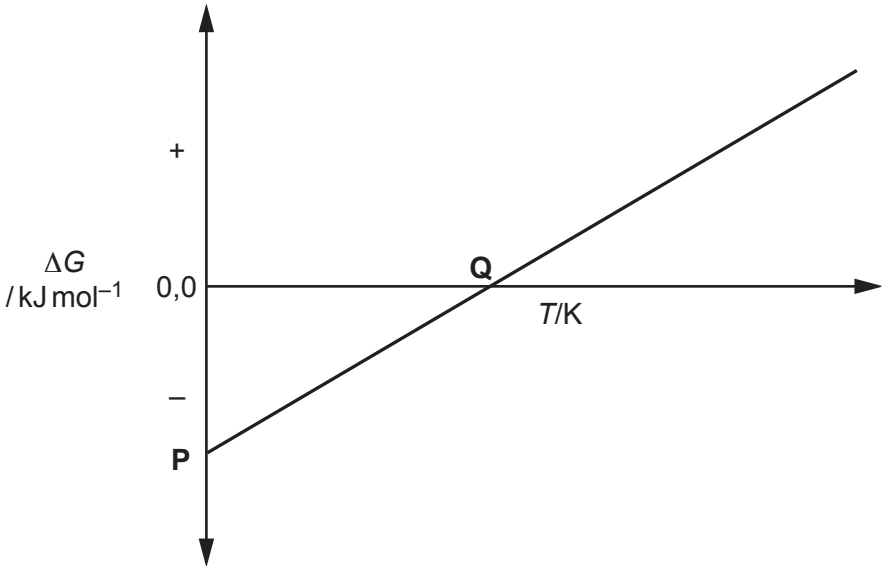
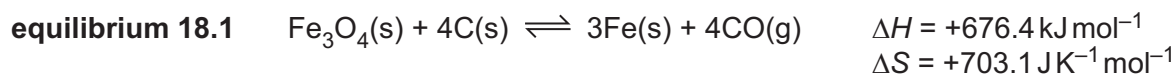


Fig. 18.1

What is significant about the gradient of the line and the values **P** and **Q** shown in **Fig. 18.1**? Explain your reasoning.

[4]

- (b) Iron can be extracted from its ore Fe_3O_4 using carbon.
Several equilibria are involved including **equilibrium 18.1**, shown below.



- (i) Why is **equilibrium 18.1** a *heterogeneous* equilibrium?

.....
 [1]

- (ii) Write the expression for K_p for **equilibrium 18.1**.

[1]

- (iii) The forward reaction in **equilibrium 18.1** is only feasible at high temperatures.

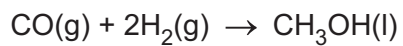
- Show that the forward reaction is **not** feasible at 25°C .
- Calculate the minimum temperature, in K, for the forward reaction to be feasible.

minimum temperature = ... K [3]

- 3 The table below shows standard entropies, $S^\ominus$.

Substance	CO(g)	H ₂ (g)	CH ₃ OH(l)
$S^\ominus / \text{J mol}^{-1} \text{K}^{-1}$	197.6	130.6	239.7

What is the entropy change, $\Delta S^\ominus$, in $\text{J mol}^{-1} \text{K}^{-1}$, for the following reaction?



- A** -219.1
B -88.5
C +88.5
D +219.1

Your answer

[1]

4 This question is about enthalpy changes.

- (a) **Table 16.1** shows enthalpy changes that can be used to determine the enthalpy change of hydration of fluoride ions, F^- .

Enthalpy change	Energy/ kJ mol^{-1}
Hydration of Ca^{2+}	-1609
Solution of CaF_2	+13
Lattice enthalpy of CaF_2	-2630

Table 16.1

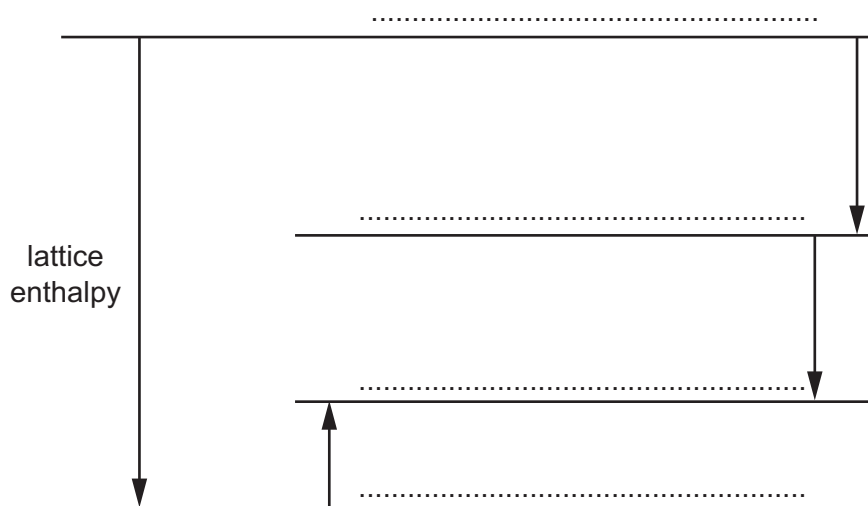
- (i) Explain what is meant by the term *enthalpy change of hydration*.

.....

 [2]

- (ii) The enthalpy change of hydration of F^- can be determined using the enthalpy changes in **Table 16.1** and the incomplete energy cycle below.

On the dotted lines, add the species present, including state symbols.



[4]

- (iii) Calculate the enthalpy change of hydration of fluoride ions, F^- .

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

- (iv) Predict how the enthalpy changes of hydration of F^- and Cl^- would differ.

Explain your answer.

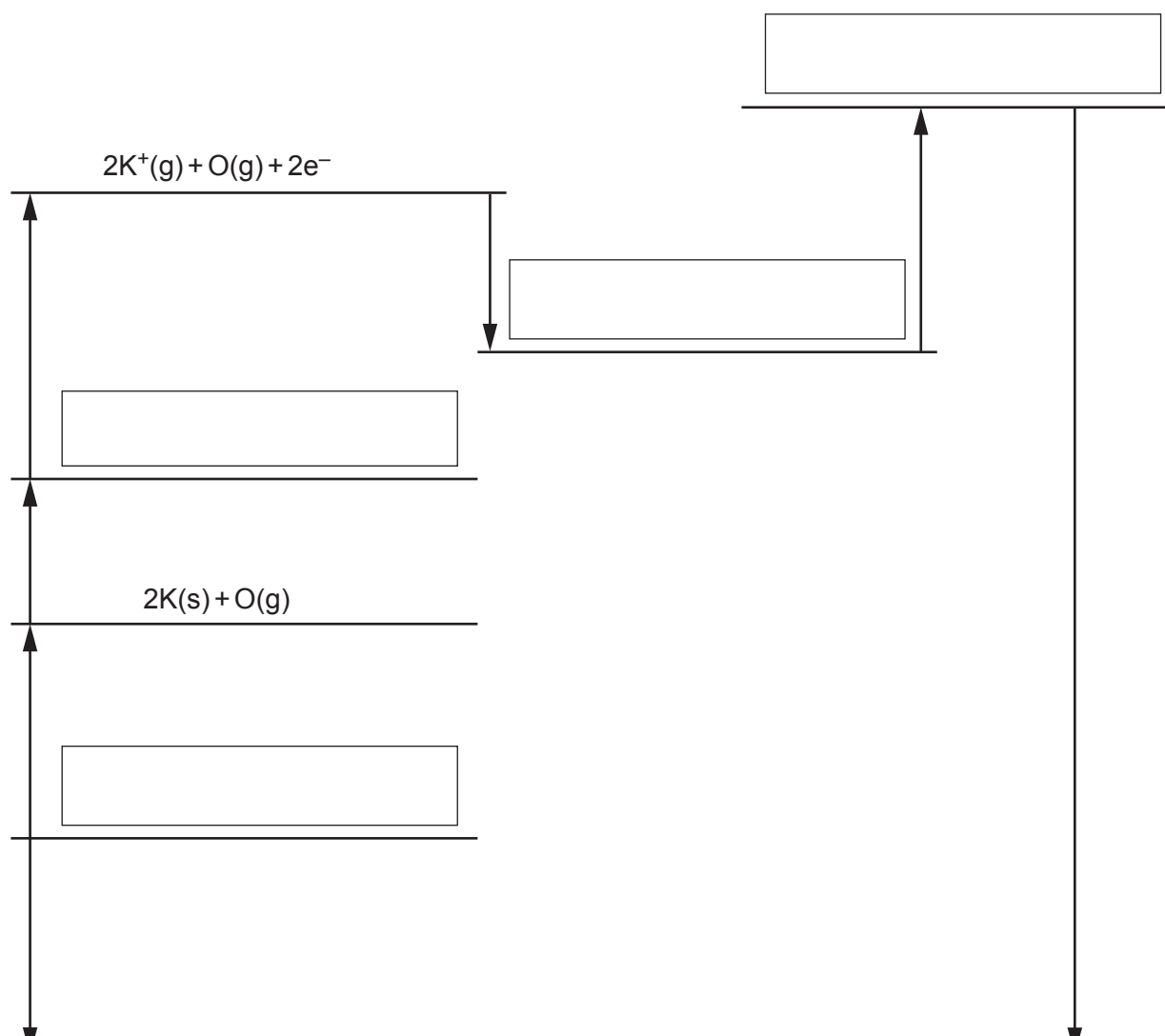
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..... [2]

- 5(a) The table below shows enthalpy changes involving potassium, oxygen and potassium oxide, K_2O .

	Enthalpy change / kJ mol^{-1}
formation of potassium oxide	-363
1st electron affinity of oxygen	-141
2nd electron affinity of oxygen	+790
1st ionisation energy of potassium	+419
atomisation of oxygen	+249
atomisation of potassium	+89

- (i) The incomplete Born–Haber cycle below can be used to determine the lattice enthalpy of potassium oxide.

In the boxes, complete the species present in the cycle.
Include state symbols for the species.



- (ii) Calculate the lattice enthalpy of potassium oxide.

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

- (b) A similar Born–Haber cycle to potassium oxide in (d) can be constructed for sodium oxide.

- (i) The first ionisation energy of sodium is more endothermic than that of potassium.

Explain why.

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..... [2]

- (ii) The lattice enthalpy of sodium oxide is more exothermic than that of potassium oxide.

Explain why.

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..... [2]

- 6 Sulfuric acid is an important chemical used to make detergents, fertilisers and dyes. It is manufactured in a multi-step process.

- (a) In the first step of the manufacture of sulfuric acid, sulfur dioxide, SO_2 , can be made from the combustion of hydrogen sulfide, H_2S , shown in **Reaction 1**.



- (i) Explain why the enthalpy change for **Reaction 1** has a negative value.

Use ideas about enthalpy changes associated with bond breaking and bond making.

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 [1]

- (ii) Some standard entropy values are given below.

Substance	$\text{H}_2\text{S}(\text{g})$	$\text{O}_2(\text{g})$	$\text{SO}_2(\text{g})$	$\text{H}_2\text{O}(\text{l})$
$S^\circ / \text{JK}^{-1} \text{mol}^{-1}$	206	205	248	70

Using calculations, explain whether **Reaction 1** is feasible at 20°C .

Calculations

Explanation for feasible or non feasible
 [4]

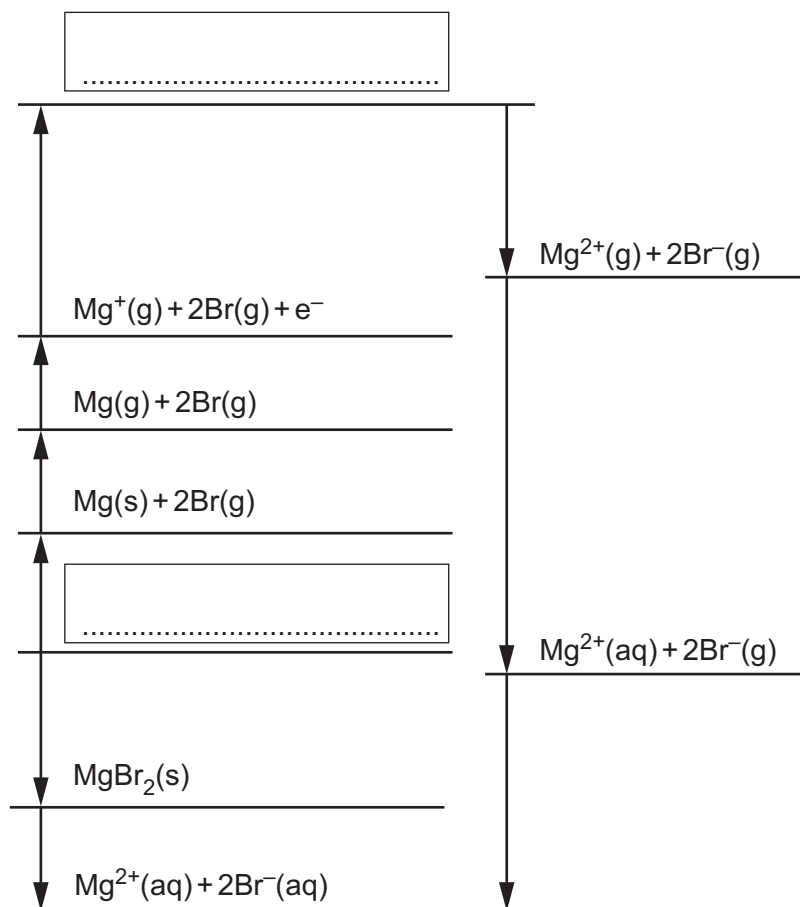
- 7(a) The enthalpy change of hydration of bromide ions can be determined using the enthalpy changes in **Table 16.2**.

Enthalpy change	Energy / kJ mol^{-1}
1st ionisation energy of magnesium	+736
2nd ionisation energy of magnesium	+1450
atomisation of bromine	+112
atomisation of magnesium	+148
electron affinity of bromine	-325
formation of magnesium bromide	-525
hydration of bromide ion	to be calculated
hydration of magnesium ion	-1926
solution of magnesium bromide	-186

Table 16.2

- (i) An incomplete energy cycle based on **Table 16.2** is shown below.

On the dotted lines, add the species present, including state symbols.

**[2]**

- (ii) Using your completed energy cycle in **16(d)(i)**, calculate the enthalpy change of hydration of bromide ions.

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

- (iii) Write the equation for the lattice enthalpy of magnesium bromide and calculate the lattice enthalpy of magnesium bromide.

Equation

Calculation

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

- 8 Methanol, CH₃OH, can be made industrially by the reaction of carbon monoxide with hydrogen, as shown in **equilibrium 1**.



- (a) Standard entropy values are given below.

Substance	CO(g)	H ₂ (g)	CH ₃ OH(g)
$S^\circ / \text{J K}^{-1} \text{ mol}^{-1}$	198	131	238

A chemist proposed producing methanol at 525 K using **equilibrium 1**.

Explain, with a calculation, whether the production of methanol is feasible at 525 K.

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..... [5]

- 9 The feasibility of a chemical reaction depends on the temperature and the signs of ΔH and ΔS .

Which row has signs for ΔH and ΔS for a reaction that is feasible at high temperatures but is **not** feasible at low temperatures?

	ΔH	ΔS
A	+	+
B	+	–
C	–	+
D	–	–

Your answer

[1]

- 10(a)** The enthalpy change of solution for magnesium fluoride, MgF_2 , can be determined indirectly using an energy cycle based on the enthalpy changes below.

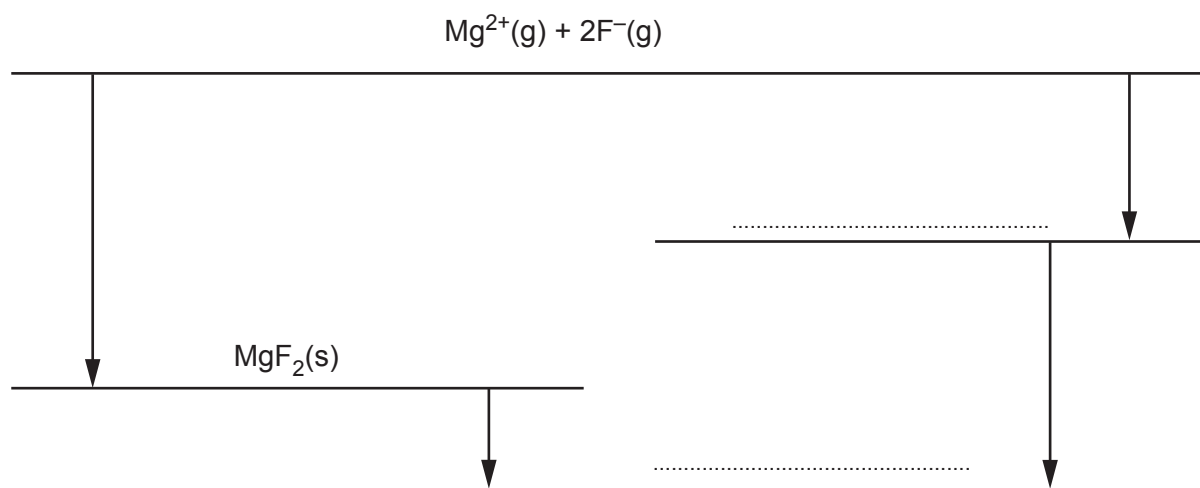
Enthalpy change	Energy / kJ mol^{-1}
Lattice enthalpy of magnesium fluoride	–2926
Hydration of magnesium ions	–1920
Hydration of fluoride ions	–506

- (i) Explain what is meant by **enthalpy change of solution**.

.....

 [1]

- (ii) On the dotted lines, add the species present, including state symbols.



[2]

- (iii) Calculate the enthalpy change of solution of MgF_2 .

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

- (iv) The enthalpy changes of solution of the magnesium halides show a trend from MgF_2 to MgI_2 .

Explain why it is difficult to predict whether the enthalpy change of solution becomes more exothermic or less exothermic down the group from MgF_2 to MgI_2 .

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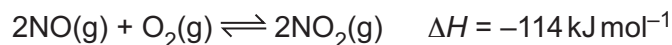
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..... [4]

- 11 This question is about chemical equilibrium.

Nitrogen monoxide, NO, and oxygen, O₂, react to form nitrogen dioxide, NO₂, in the reversible reaction shown in **Equilibrium 20.1**.

Equilibrium 20.1



$$\Delta S = -147 \text{ J mol}^{-1} \text{ K}^{-1}$$

- (a) A dynamic equilibrium exists in a closed system.

State **one** other feature of a dynamic equilibrium.

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 [1]

- (b) (i) Show that the formation of NO₂ in **Equilibrium 20.1** is feasible at 25 °C.

[2]

- (ii) Determine the maximum temperature, in K, for feasibility.

Give your answer to an **appropriate** number of significant figures.

maximum temperature = ... K [1]

12 For the condensation of ammonia gas, what are the signs of ΔH and ΔS ?

A ΔH -ve ΔS -ve

B ΔH -ve ΔS +ve

C ΔH +ve ΔS +ve

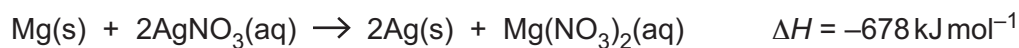
D ΔH +ve ΔS -ve

Your answer

[1]

13 This question is about energy changes.

(a) Magnesium reacts with aqueous silver nitrate, $\text{AgNO}_3(\text{aq})$ as shown below.



A student adds an excess of magnesium to 100.0 cm^3 of $0.400 \text{ mol dm}^{-3}$ $\text{AgNO}_3(\text{aq})$.
The initial temperature is 20.0°C .

(i) Determine the maximum temperature reached in this reaction.
Give your answer to **3** significant figures.

Assume that the specific heat capacity and density of the solution are the same as for water, and that there are no heat losses.

maximum temperature reached = ... $^\circ\text{C}$ **[4]**

- (ii) The student wants to repeat the experiment, but there is not enough $\text{AgNO}_3(\text{aq})$ left to use another 100.0 cm^3 portion.

The student decides to modify the method by adding an excess of magnesium to 50.0 cm^3 of $0.400\text{ mol dm}^{-3}\text{AgNO}_3(\text{aq})$.

Predict, with reasons, how this modification would affect the maximum temperature reached. Assume that there are no heat losses.

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..... [1]

- (b) Nitric acid is manufactured from ammonia in a multi-stage process.
The equation for the first stage in this process is shown in **Reaction 17.1**.



Some standard enthalpy changes of formation are shown in the table.

Compound	$\Delta_f H^\ominus / \text{kJ mol}^{-1}$
$\text{NH}_3(\text{g})$	-46
$\text{H}_2\text{O}(\text{l})$	-286

- (i) Explain the term **enthalpy change of formation**.

.....

 [1]

- (ii) Calculate the standard enthalpy change of formation, $\Delta_f H^\ominus$, of $\text{NO}(\text{g})$.

$\Delta_f H^\ominus$ of $\text{NO}(\text{g}) = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

- (c) Carbon disulfide, CS_2 , reacts with dinitrogen oxide, N_2O , as shown in **Reaction 17.2**.



Standard entropies, $S^\ominus$, are shown in the table.

Substance	$\text{CS}_2(\text{l})$	$\text{N}_2\text{O}(\text{g})$	$\text{S}_8(\text{s})$	$\text{CO}_2(\text{g})$	$\text{N}_2(\text{g})$
$S^\ominus/\text{JK}^{-1}\text{mol}^{-1}$	151	220	256	214	192

- (i) Explain the term **entropy**.

.....

 [1]

- (ii) The free energy change, ΔG , of **Reaction 17.2** is $-2672 \text{ kJ mol}^{-1}$ at 25°C .

Calculate the enthalpy change, ΔH , of **Reaction 17.2**, in kJ mol^{-1} .

$\Delta H = \dots\dots\dots \text{kJ mol}^{-1}$ [3]

- (iii) A student concludes that **Reaction 17.2** is feasible at all temperatures.

Explain whether the student is correct or not.

.....

 [2]

- 14 The table below shows standard entropies, $S^\ominus$.

Substance	SO ₂ (g)	O ₂ (g)	SO ₃ (l)
$S^\ominus/\text{JK}^{-1}\text{mol}^{-1}$	248	204	96

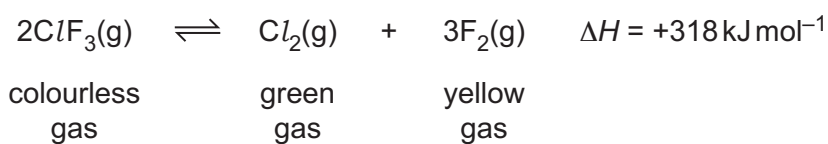
What is the standard entropy change, $\Delta S^\ominus$, in JKmol^{-1} , for the formation of 1 mol of SO₃(l) from SO₂(g) and O₂(g)?

- A -508
B -254
C +254
D +508

Your answer

[1]

- 15** Chlorine trifluoride can be decomposed into its elements forming the equilibrium mixture below.



Which statement(s) is/are correct?

- 1** The decomposition is a redox reaction.
 - 2** When the equilibrium mixture is cooled, the colour fades.
 - 3** The decomposition has a negative entropy change.
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- A** 1, 2 and 3
 - B** Only 1 and 2
 - C** Only 2 and 3
 - D** Only 1

Your answer

[1]

16 This question is about energy changes.

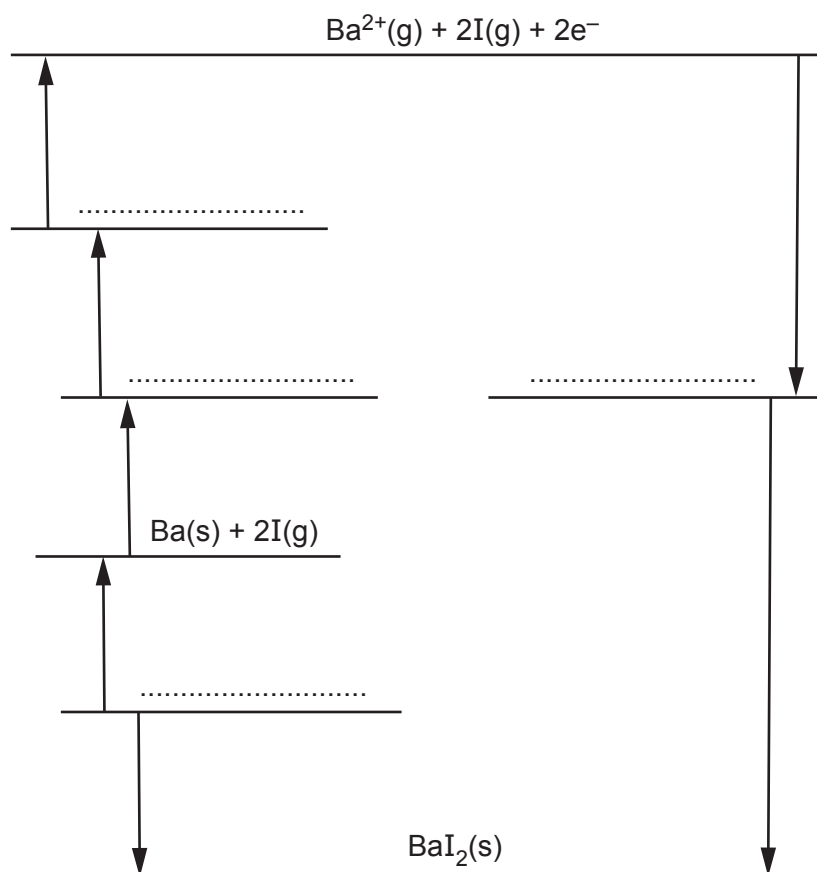
(a) Lattice enthalpies can be determined indirectly using Born-Haber cycles.

The table below shows the energy changes that are needed to determine the lattice enthalpy of barium iodide, BaI_2 .

Energy term	Energy change / kJ mol^{-1}
formation of barium iodide	-602
1st electron affinity of iodine	-296
1st ionisation energy of barium	+503
2nd ionisation energy of barium	+965
atomisation of iodine	+107
atomisation of barium	+180

(i) The diagram below shows an incomplete Born-Haber cycle that can be used to calculate the lattice enthalpy of barium iodide.

On the dotted lines, add the species present, including state symbols.



[4]

(ii) Calculate the lattice enthalpy of barium iodide.

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

17 This question is about enthalpy changes of reactions involving hydrocarbons.

(a) A student determines the enthalpy change of combustion, $\Delta_c H$, of heptane, C_7H_{16} , using the method outlined below.

- Add 150 g of water to a beaker and measure its temperature.
- Weigh a spirit burner containing heptane and use it to heat the water.
- Extinguish the flame and record the maximum temperature reached by the water.
- Reweigh the spirit burner.

The temperature of the water increased by 10.5°C .

The spirit burner decreased in mass by 0.133 g.

Use the student's results to determine the enthalpy change of combustion of heptane, $\Delta_c H (C_7H_{16})$, in kJ mol^{-1} .

$\Delta_c H (C_7H_{16}) = \dots\dots\dots \text{kJ mol}^{-1}$ **[3]**

- (b) Nonane, C_9H_{20} , can be broken down by heat to form pentane, C_5H_{12} , and ethene, C_2H_4 .



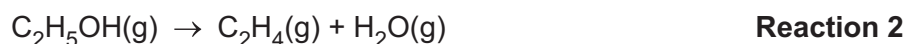
The enthalpy changes of combustion of $C_9H_{20}(g)$ and $C_2H_4(g)$ are shown in the table below.

Hydrocarbon	$\Delta_c H / \text{kJ mol}^{-1}$
$C_9H_{20}(g)$	-6171
$C_2H_4(g)$	-1411

Use ΔH in **Reaction 1** and the enthalpy changes of combustion in the table to determine the enthalpy change of combustion of $C_5H_{12}(g)$.

$$\Delta_c H (C_5H_{12}(g)) = \dots\dots\dots \text{ kJ mol}^{-1} \quad [2]$$

- (c) Ethene can be produced from ethanol, as shown in **Reaction 2** below.



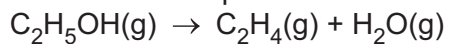
- (i) Predict the sign of the entropy change, ΔS , for **Reaction 2**.

Explain your reasoning.

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 [1]

(ii) **Reaction 2** is repeated:



Reaction 2

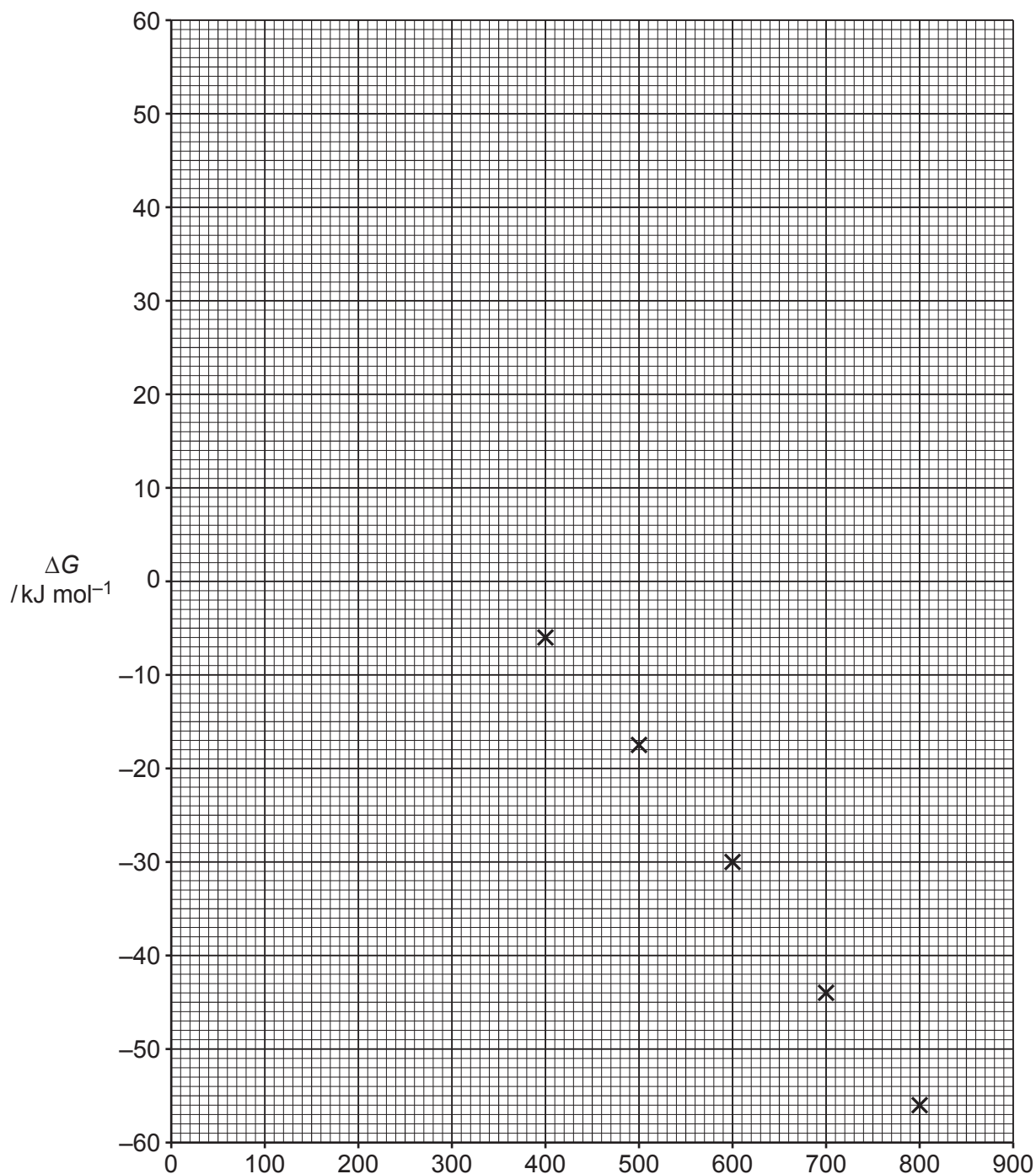
The Gibbs equation is shown below.

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

The enthalpy change, ΔH , and the entropy change, ΔS , can be assumed to be constant at different temperatures.

Fig. 18.1 shows values of the free energy change, ΔG , in kJ mol^{-1} , at different temperatures, T , in K, for **Reaction 2**.

Fig. 18.1



Use the graph in **Fig. 18.1** to answer the following:

- Draw the best-fit line on the graph in **Fig. 18.1**.
- Determine ΔS , in $\text{JK}^{-1} \text{mol}^{-1}$, for **Reaction 2**.
- Determine the minimum temperature, T , at which the reaction is feasible.
- Determine ΔH for **Reaction 2**.

$\Delta S = \dots\dots\dots \text{JK}^{-1} \text{mol}^{-1}$

minimum $T = \dots\dots\dots \text{K}$

$\Delta H = \dots\dots\dots \text{kJmol}^{-1}$

18 This question is about energy changes.

Hydrogen peroxide decomposes as shown in **Reaction 16.1**.



(a) The table shows enthalpy changes of formation and entropies.

	$\Delta H_f^\circ / \text{kJ mol}^{-1}$	$S^\circ / \text{J K}^{-1} \text{mol}^{-1}$
$\text{H}_2\text{O}_2(\text{l})$	-188	110
$\text{H}_2\text{O}(\text{l})$	-286	70.0
$\text{O}_2(\text{g})$	0	205

(i) Calculate the free-energy change, ΔG , in kJ mol^{-1} , of **Reaction 16.1** at 25°C .

Give your answer to 3 significant figures.

$\Delta G = \dots\dots\dots \text{kJ mol}^{-1}$ [4]

(ii) The decomposition of hydrogen peroxide shown in **Reaction 16.1** is feasible.

Suggest why **Reaction 16.1** does **not** take place at 25°C despite being feasible.

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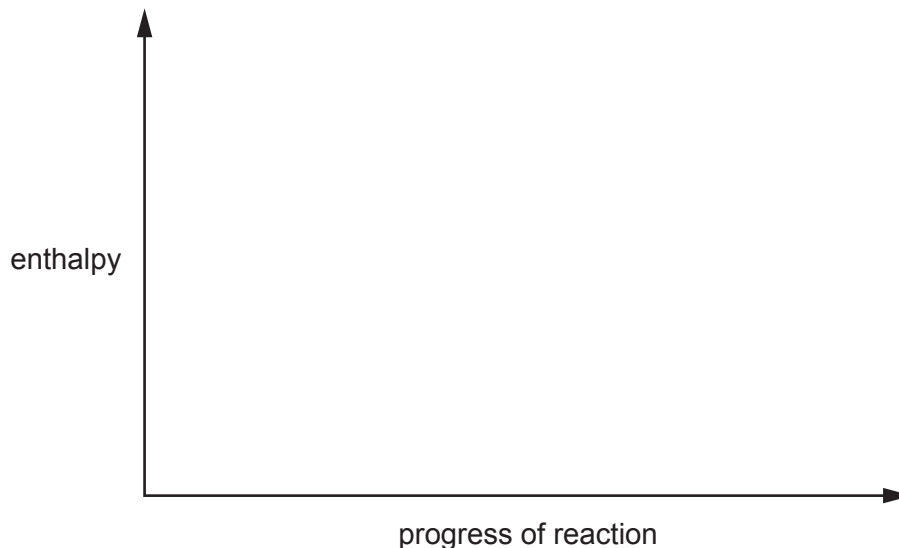
..... [1]

- (b) The rate of decomposition of hydrogen peroxide shown in **Reaction 16.1** can be increased by adding a small amount of powdered manganese(IV) oxide, MnO_2 .

The MnO_2 acts as a catalyst.

- (i) Complete the enthalpy profile diagram for **Reaction 16.1** using formulae for the reactants and products.

- Use E_a to label the activation energy **without** MnO_2 .
- Use E_c to label the activation energy **with** MnO_2 .
- Use ΔH to label the enthalpy change of reaction.



[3]

- (ii) Explain why MnO_2 is described as a **heterogeneous** catalyst for this reaction.

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 [1]

- (iii) Mn_3O_4 is a compound in which Mn has two different oxidation states. The two oxidation states are different from the Mn in MnO_2 .

Suggest the two oxidation states of manganese in Mn_3O_4 .

..... [1]

- (c) Manganese(II) oxide, MnO , has a giant ionic lattice structure.

The table shows the enthalpy changes that are needed to determine the lattice enthalpy of MnO .

	enthalpy change / kJ mol^{-1}
atomisation of manganese	+281
atomisation of oxygen	+249
first ionisation energy of manganese	+717
second ionisation energy of manganese	+1509
first electron affinity of oxygen	-141
second electron affinity of oxygen	+798
formation of manganese(II) oxide	-385

- (i) Define the term **lattice enthalpy**.

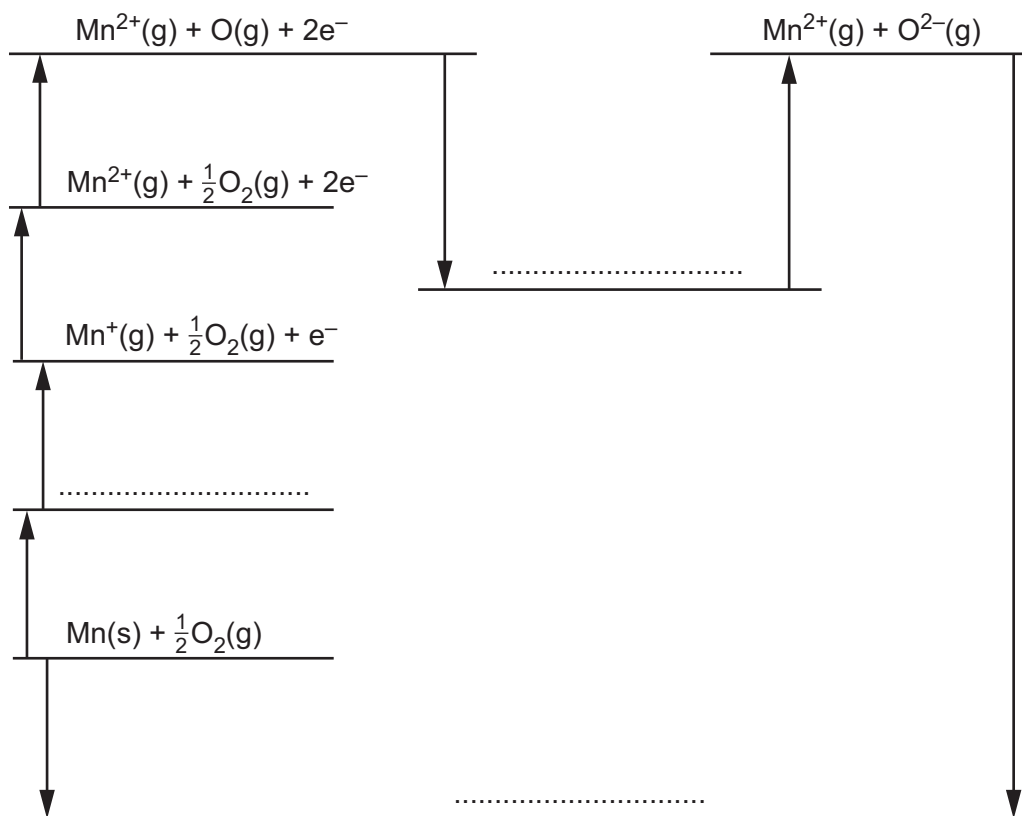
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- (ii) The diagram shows an incomplete Born-Haber cycle that can be used to determine the lattice enthalpy of MnO.



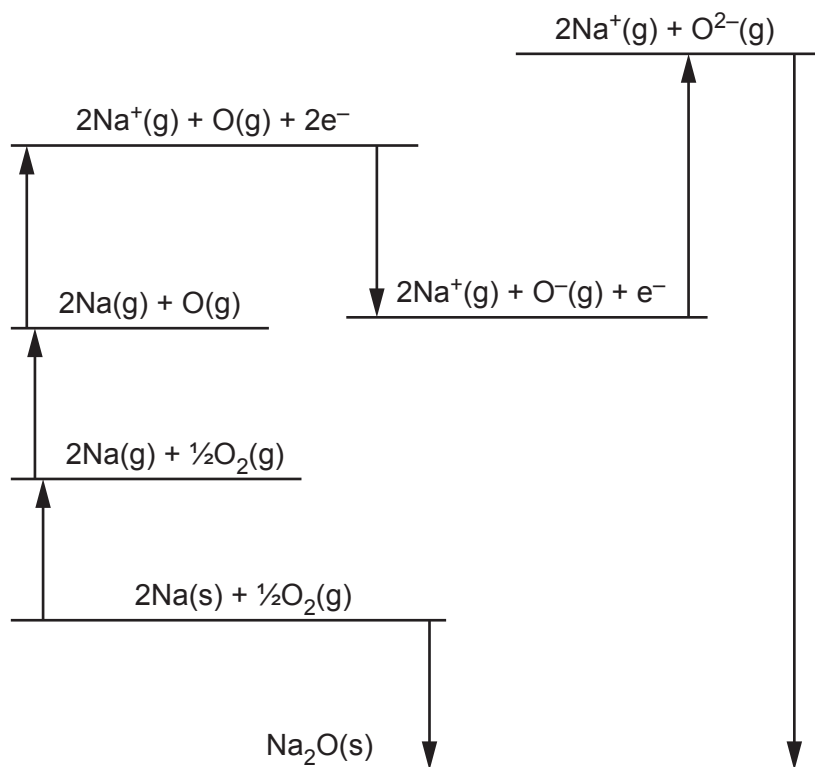
Complete the diagram by adding the species present on the dotted lines, include state symbols.

[3]

- (iii) Calculate the lattice enthalpy of MnO.

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

- 19 The diagram shows the Born-Haber cycle for sodium oxide, Na_2O .



Why is the value for the second electron affinity of oxygen positive?

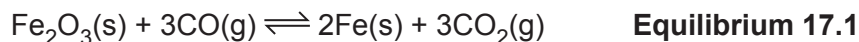
- A** An electron is added to a high energy orbital.
- B** An electron is added to a negative ion.
- C** An electron is removed from a gaseous atom.
- D** An electron is removed from a positive ion.

Your answer

[1]

20 This question is about carbon monoxide and carbon as reducing agents in industry.

(a) Iron(III) oxide, Fe_2O_3 , can be reduced by carbon monoxide to form iron.



Some standard enthalpy change of formation, $\Delta_f H^\ominus$, and standard entropy, $S^\ominus$, values are shown in the table.

Substance	$\text{Fe}_2\text{O}_3(\text{s})$	$\text{CO}(\text{g})$	$\text{Fe}(\text{s})$	$\text{CO}_2(\text{g})$
$\Delta_f H^\ominus / \text{kJ mol}^{-1}$	−824	−111	0	−394
$S^\ominus / \text{JK}^{-1} \text{mol}^{-1}$	87	198	To calculate in part (a)(iv)	214

(i) Explain why the standard enthalpy change of formation of Fe has a value of 0 kJ mol^{-1} .

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 [1]

(ii) Calculate the enthalpy change, $\Delta_r H$, for the forward reaction in **Equilibrium 17.1**.

$$\Delta H = \dots\dots\dots \text{kJ mol}^{-1} \quad [2]$$

(iii) Explain why the standard entropy of CO_2 is greater than the standard entropy of Fe_2O_3 .

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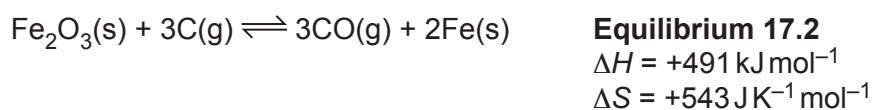
 [2]

- (iv) The standard entropy change, $\Delta S^\ominus$, for the forward reaction in **Equilibrium 17.1** has a value of $+15 \text{ JK}^{-1} \text{ mol}^{-1}$.

Use data from the table to calculate the standard entropy, $S^\ominus$, of Fe.

$$S^\ominus = \dots\dots\dots \text{ JK}^{-1} \text{ mol}^{-1} \quad [2]$$

- (b) Carbon can also reduce $\text{Fe}_2\text{O}_3(\text{s})$ to iron as shown in **Equilibrium 17.2**.



The forward reaction is only feasible at high temperatures.

- Show that the forward reaction is **not** feasible at 100°C .
- Calculate the minimum temperature, in K, for the forward reaction to be feasible.

Give the minimum temperature to the nearest whole number.

$$T = \dots\dots\dots \text{ K} \quad [3]$$