NAME DATE STARTED TARGET END DATE

GCE A LEVEL — CHEMISTRY UNIT 3 QUESTION PACK

1095-01 (Legacy CH5) · New spec Unit 3 Topic 5 · A2 unit, first sat 2017, 80 marks, 1h 45min paper

REVISE.wales

CHEMISTRY — UNIT 3 · Thermodynamics — Born-Haber, Entropy & Gibbs

Topics 3.6 & 3.7 — Lattice energies, Born-Haber cycles, enthalpies of solution and hydration, entropy and Gibbs free energy

Constructing Born-Haber cycles to find lattice energy, predicting trends from ionic polarisation (Fajans' rules), explaining enthalpies of solution / hydration, and using ΔG = ΔH – TΔS to predict the feasibility of a reaction.

Legacy 2008 specification

Estimated time for entire question pack: ~1 h 41 min

Derived from the legacy CH5 paper's pace of ~1.3 min/mark, padded for long-prose and calculation answers (63 marks over 5 questions).

*You are advised to **not** attempt to complete all of this in one sitting.*

ABOUT THIS QUESTION PACK

This is a **comprehensive practice question pack**, not a single mock paper. It contains every question from the legacy WJEC CH5 papers (2008 modular spec, Jun 2010 – Jun 2016) that maps onto the new-spec A2 Unit 3 Topic 3.6 & 3.7.

Questions are ordered by source paper date.

INSTRUCTIONS

Use black ink or black ball-point pen. Show all working — quality of written communication will affect marks. A calculator is allowed. You will need the WJEC Periodic Table / Data Booklet.

For Examiner's use only

Q	Source	Max	Mark
1	Jun 10 Q1	12	
2	Jun 11 Q2	11	
3	Jun 12 Q2	11	
4	Jun 13 Q5	20	
5	Jun 15 Q2	9	
Total		**63**	

Thermodynamics — Born-Haber, Entropy & Gibbs — what the new spec asks

WJEC GCE A Level Chemistry (from 2015) · Unit 3: Physical & Inorganic Chemistry · Topic 3.6 & 3.7.

Lattice energy

- ΔH_{latt}: enthalpy when 1 mol of solid ionic compound formed from gaseous ions.
- Always exothermic (negative).
- More negative when ion charges ↑ or ionic radii ↓.
- Theoretical lattice energy assumes purely ionic; covalent character makes experimental more exothermic.

Born-Haber cycle

- Cycle relates ΔH_f to atomisation, ionisation, electron affinity, lattice.
- $\Delta H_f = \Sigma(\Delta H_{at}) + \Sigma(IE) + \Sigma(EA) + \Delta H_{latt}$.
- All enthalpies are 'per mole' and standard (298 K, 1 atm).
- Use Hess's law to find the unknown step.

Solution & hydration

- $\Delta H_{sol} = \Delta H_{latt} + \Sigma\Delta H_{hyd}$ (cations + anions).
- ΔH_{hyd} exothermic (water dipole attracted to ion).
- Small, highly charged ions ⇒ more exothermic hydration.
- Soluble ⇒ $|\Delta H_{hyd}| \geq |\Delta H_{latt}|$.

Fajans' rules & polarisation

- Small, highly charged cation polarises anion electron cloud.
- Large, highly charged anion is easily polarised.
- ⇒ more covalent character; experimental lattice E differs from ionic model.
- $AlCl_3$ is more covalent than NaCl; AgI more covalent than AgF.

Entropy

- ΔS = total disorder change. Gas > liquid > solid.
- Solids dissolving: usually $\Delta S > 0$.
- More moles of gas on products side ⇒ $\Delta S > 0$.
- Universe: $\Delta S_{total} = \Delta S_{system} + \Delta S_{surroundings}$.

Gibbs free energy & feasibility

- $\Delta G = \Delta H - T\Delta S$ (T in K, ΔS in J K^{-1} mol^{-1} ⇒ convert!).
- $\Delta G < 0$ ⇒ thermodynamically feasible.
- Endothermic + $\Delta S > 0$: feasible if $T > \Delta H/\Delta S$.
- Exothermic + $\Delta S < 0$: feasible if $T < \Delta H/\Delta S$.

Thermodynamics — Born-Haber, Entropy & Gibbs in one page

Quick-reference notes — revisit before each question.

Born-Haber cycle (NaCl)

Na(s) → Na(g): $+\Delta H_{at}$

Na(g) → Na^+(g): $+IE_1$

½Cl_2(g) → Cl(g): $+\Delta H_{at}$

Cl(g) → Cl^-(g): $-EA_1$

Na^+(g) + Cl^-(g) → NaCl(s): $+\Delta H_{latt}$

Sum = ΔH_f(NaCl).

Lattice energy trend

ΔH_{latt} ↑ (more −ve) with q^+q^-/r ⇒ MgO > NaCl. Smaller / more highly-charged ions ⇒ more exothermic.

$\Delta H_{sol} = \Delta H_{latt} + \Delta H_{hyd}$

For NaCl: ΔH_{latt} = +787, $\Delta H_{hyd}(Na^+)$ = −406, $\Delta H_{hyd}(Cl^-)$ = −378 ⇒ ΔH_{sol} = +3 kJ mol^{-1} (slight endothermic).

Fajans' rules

More covalent when (a) small high-charge cation, (b) large high-charge anion, (c) cation with non-noble e^- config (e.g. d^{10} Ag^+).

ΔS signs

+ve: more gas moles; solid → liquid → gas; mixing.
−ve: gas moles ↓; gas → liquid; gas → solid.

Feasibility temp

Endothermic + ΔS > 0: spontaneous above $T = \Delta H/\Delta S$.
Exothermic + ΔS < 0: spontaneous below that T. Remember: ΔS must be in kJ K^{-1} mol^{-1}!

Questions begin overleaf.

SECTION A

Answer **all** *the questions in the spaces provided.*

1. *(a)* Magnesium carbonate decomposes on heating.

$$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$$

(i) Given the enthalpy change of formation, $\Delta H_f^{\ominus}$, values below, calculate the enthalpy change, $\Delta H^{\ominus}$, for the decomposition of magnesium carbonate. [1]

Species	Enthalpy change of formation $\Delta H_f^{\ominus}$ / kJ mol^{-1}
$CO_2(g)$	–393.5
$MgCO_3(s)$	–1095.8
$MgO(s)$	–601.7

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(ii) The entropy change, $\Delta S^{\ominus}$, for the decomposition is 174.8 J mol^{-1}K^{-1}. Explain why there is an increase in entropy for this reaction. [1]

....

....

(iii) Convert the value of $\Delta S^{\ominus}$ into units of kJ mol^{-1}K^{-1}. [1]

....

....

(iv) Using your answers to *(a)*(i) and (iii), determine, in degrees K, the temperature above which magnesium carbonate would decompose spontaneously. [3]

....

....

....

(b) The solution of ionic compounds such as magnesium carbonate or sodium carbonate in water at 20 °C (room temperature) can be represented by the equations

$$MgCO_3(s) \rightleftharpoons Mg^{2+}(aq) + CO_3^{2-}(aq)$$

$$Na_2CO_3(s) \rightleftharpoons 2Na^{+}(aq) + CO_3^{2-}(aq)$$

Use the free energy change, ΔG, values in the table to comment on the solubilities of magnesium carbonate and sodium carbonate in water. [2]

Solution	Free Energy Change ΔG / kJ mol^{-1}
$MgCO_3(s) \rightleftharpoons Mg^{2+}(aq) + CO_3^{2-}(aq)$	+28.2
$Na_2CO_3(s) \rightleftharpoons 2Na^{+}(aq) + CO_3^{2-}(aq)$	–4.3

(c) As solids do not affect the position of equilibrium, for the solution equilibrium

$$MgCO_3(s) \rightleftharpoons Mg^{2+}(aq) + CO_3^{2-}(aq)$$

the simplest expression for the equilibrium constant, K_c, can be written

$$K_c = [Mg^{2+}(aq)][CO_3^{2-}(aq)]$$

(i) Given that the solubility of $MgCO_3$ at 20 °C is 3.16×10^{-3} mol dm^{-3}, state the molar concentrations of magnesium ions, Mg^{2+}(aq), and carbonate ions, CO_3^{2-}(aq), in a saturated $MgCO_3$ solution. [1]

....

....

(ii) Hence calculate the value of K_c at 20 °C. [1]

....

....

(iii) Giving your reasons, state whether the value of K_c is consistent with the value of the free energy change, ΔG, given for this reaction in *(b)*. [1]

....

....

(iv) By applying Le Chatelier's Principle to the chemical equation above, and giving your reasons, state the effect on the solubility of magnesium carbonate of adding sodium carbonate to the solution. [1]

....

....

Total [12]

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2. Fuel cells have been proposed as an alternative method of providing energy for vehicles. These use chemical reactions within electrochemical systems to generate electricity.

(a) A typical fuel cell uses hydrogen as a fuel and reacts this with oxygen. The two half-equations for the processes occurring at the electrodes are given in the table below.

Half-equation	$E^{\ominus}$ / V
$2H^+ + 2e^- \rightleftharpoons H_2$	0.00
$\frac{1}{2}O_2 + 2H^+ + 2e^- \rightleftharpoons H_2O$	1.23

(i) Write an equation for the overall reaction occurring. [1]

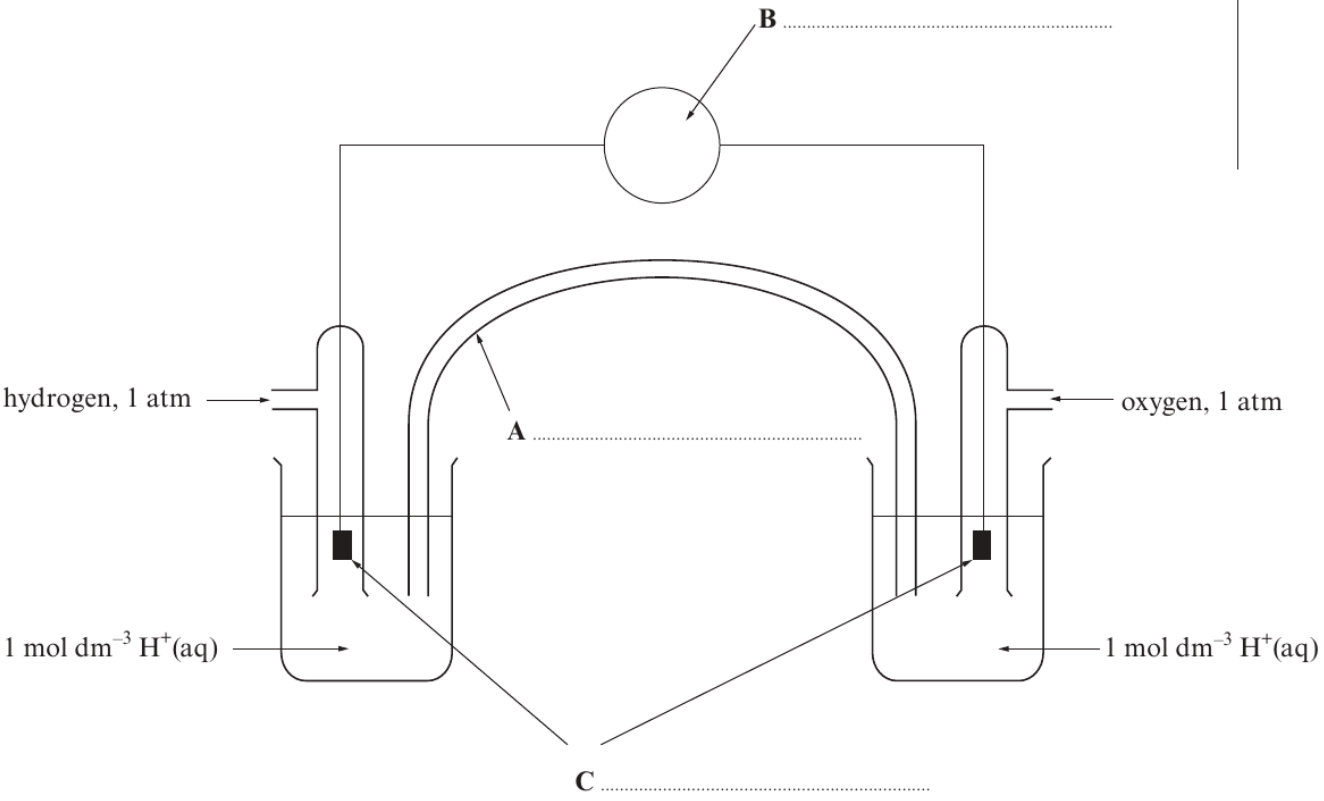
....

(ii) Give **one** benefit of the use of fuel cells as a replacement for traditional vehicle energy sources. [1]

....

....

(iii) The same reaction as above can be undertaken in a traditional electrochemical cell, such as the one below. Name the parts labelled **A**-**C**. [3]



(b) A different fuel for use in fuel cells is methanol, CH_3OH, which would undergo the following reaction with oxygen.

$$CH_3OH(l) + 1\tfrac{1}{2}O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

Compound	Standard enthalpy change of formation, $\Delta H_f^{\ominus}$ / kJ mol^{-1}
CH_3OH	–239
CO_2	–394
H_2O	–286

(i) Calculate the standard enthalpy change of combustion for methanol. [2]

(ii) The entropy change of this reaction is calculated as follows.

ΔS = (Sum of all entropies for products) – (Sum of all entropies for reactants)

$$\Delta S = 354 - 435$$

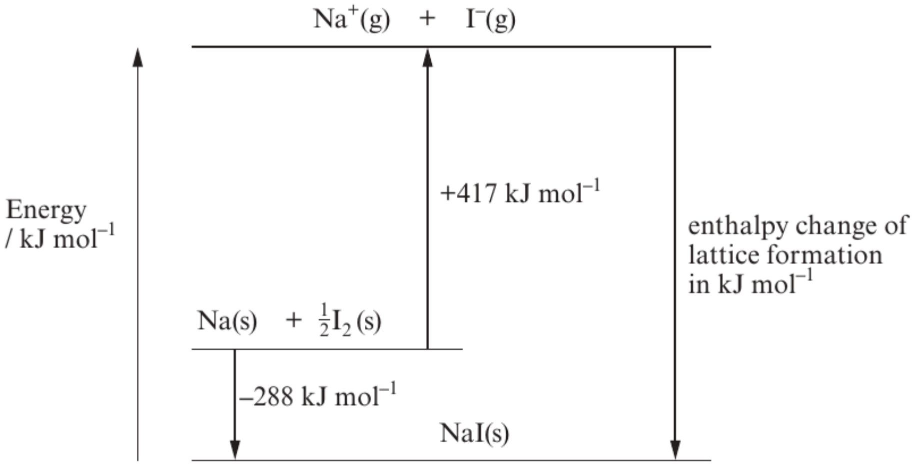
$$\Delta S = -81 \text{ J K}^{-1} \text{ mol}^{-1}$$

The reaction was repeated using gaseous methanol, $CH_3OH(g)$, in place of the liquid methanol, $CH_3OH(l)$, used above. What effect, if any, would this have on the value of the entropy change ΔS given above? Explain your answer. [2]

(iii) Use the values in parts (i) and (ii) of this question to calculate the value of the Gibbs free energy, ΔG, for this reaction at 298K and state what information this gives about the feasibility of the reaction. [2]

Total [11]

2. *(a)* The diagram shows an outline of the Born-Haber cycle for the formation of sodium iodide (NaI) from its elements.



Use the information given to calculate the enthalpy change of lattice formation (in kJ mol^{-1}) of sodium iodide. [2]

(b) Sodium iodide is very soluble in water at room temperature.

(i) Complete the sentence below using the relevant enthalpy terms.

For a compound to be very soluble in water the value of the enthalpy of

.... will be greater than the enthalpy of .... [1]

(ii) Aqueous solutions of sodium iodide become yellow in the presence of oxygen due to the slow production of iodine.
One suggested reason for this is that a low concentration of hydrogen ions in the solution produces iodine according to the equation below.

$$4H^+(aq) + 4I^-(aq) + O_2(aq) \rightleftharpoons 2I_2(aq) + 2H_2O(l)$$

Use Le Chatelier's principle to suggest a reagent that you could add, apart from water, to decrease the amount of yellow iodine present. Explain your choice. [2]

(c) Sodium chloride and sodium iodide both react with concentrated sulfuric acid to give the corresponding hydrogen halide e.g.

$$\mathrm{NaI} + \mathrm{H_2SO_4} \longrightarrow \mathrm{NaHSO_4} + \mathrm{HI}$$

However, the reaction with sodium iodide continues, giving hydrogen sulfide and iodine as two of the products. This further type of reaction does not occur when sodium chloride is used in place of sodium iodide.

(i) Describe what is **seen** when solid sodium iodide is added to concentrated sulfuric acid. [2]

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....

....

(ii) The following equations show the standard electrode potentials for the Cl_2/Cl^- and I_2/I^- systems.

$$\mathrm{Cl_2} + 2e^- \rightleftharpoons 2\mathrm{Cl}^- \qquad E^{\ominus} = +1.36\,\mathrm{V}$$

$$\mathrm{I_2} + 2e^- \rightleftharpoons 2\mathrm{I}^- \qquad E^{\ominus} = +0.54\,\mathrm{V}$$

Use these values to explain why only hydrogen iodide (represented as I^- in the equation) is able to further react with concentrated sulfuric acid in this way. [2]

....

....

....

....

(d) The reaction of chlorine with sodium hydroxide solution gives aqueous sodium chlorate(I) as one of the chlorine-containing products.

(i) Give the equation for this reaction. [1]

....

(ii) State **one** use for a solution of sodium chlorate(I). [1]

....

Total [11]

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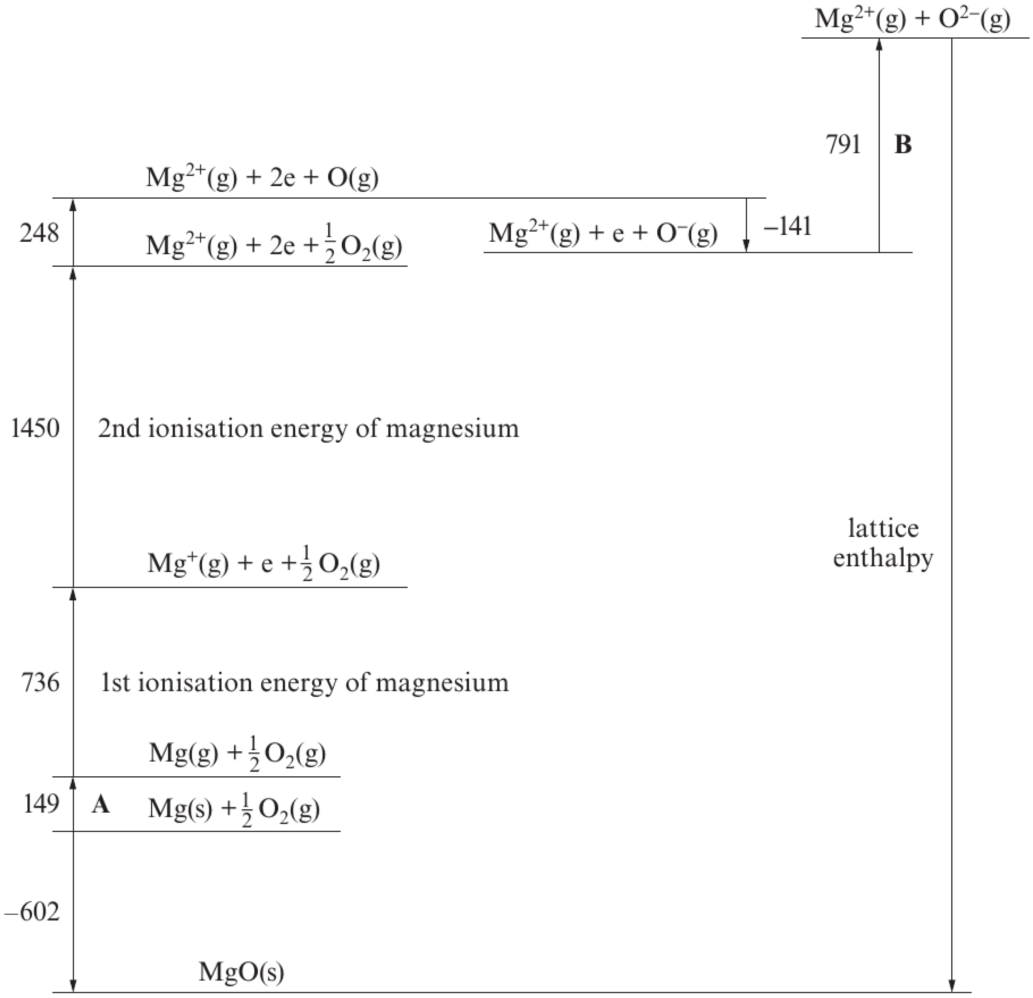
5. Magnesium oxide, MgO, is a white solid with a very high melting temperature and it is used as the refractory lining in furnaces.

(a) The following Born-Haber cycle shows the enthalpy changes involved in the formation of magnesium oxide.

All enthalpy changes are in kJ mol^{-1}. The cycle is not drawn to scale.



(i) What is the name given to the enthalpy change labelled **A**? [1]

(ii) State why the second ionisation energy of magnesium is greater than its first ionisation energy. [1]

(iii) Suggest why the second electron affinity of oxygen, labelled **B**, is positive. [1]

(iv) Calculate the value of the lattice enthalpy for magnesium oxide. [2]

(b) Many metal oxides can be reduced to the metal by carbon monoxide. The equation for the reduction of magnesium oxide is given below.

$$MgO(s) + CO(g) \longrightarrow Mg(s) + CO_2(g)$$

The conditions under which reactions will occur can be predicted using enthalpy and entropy changes. The entropies of the substances involved in this reaction are shown in the table.

Substance	$MgO(s)$	$CO(g)$	$Mg(s)$	$CO_2(g)$
Entropy / $JK^{-1}mol^{-1}$	26.9	197.7	32.7	213.7

(i) Suggest a reason why the entropies of carbon monoxide and carbon dioxide are much higher than those of magnesium and magnesium oxide. [1]

(ii) Calculate the entropy change in this reaction. [1]

(iii) The enthalpy change, ΔH, for the reduction of magnesium oxide is 318.0 kJ mol^{-1}. Calculate the minimum temperature at which this reduction could occur. [3]

(c) Magnesium oxide, MgO, lead(II) oxide, PbO, and aluminium oxide, Al_2O_3, all react with dilute acids to form aqueous ions – $Mg^{2+}(aq)$, $Pb^{2+}(aq)$ and $Al^{3+}(aq)$.

Suggest tests that would enable you to distinguish between solutions containing one of each of these ions. You should include the expected result for **each** test and are advised to record your tests and expected results in a table. [5]
QWC [2]

(d) Aluminium chloride, $AlCl_3$, can be used to produce compounds including the chloroaluminate(III) ion, $AlCl_4^-$.

(i) Draw a dot and cross diagram to show the electron arrangement in the $AlCl_4^-$ ion. You should show outer electrons only. [1]

(ii) Give **one** industrially important use in which the $AlCl_4^-$ ion is involved. State the role of the ion in this use. [2]

Total [20]

Total Section B [40]

END OF PAPER

GCE A level

1095/01-A

CHEMISTRY – CH5

Periodic Table

A.M. WEDNESDAY, 19 June 2013

2. Iron is extracted at high temperatures from the ore haematite, which contains iron(III) oxide, Fe_2O_3. The process can be summarised by the equation below.

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g) \qquad \Delta H^\theta = -23\,\text{kJ mol}^{-1}$$

Some thermodynamic data for the substances in the reaction are shown in the following table.

Substance	Standard enthalpy change of formation, ΔH^θ_f / kJ mol^{-1}	Standard entropy, S^θ / J K^{-1} mol^{-1}
$Fe_2O_3(s)$	−826	90
Fe(s)	0	27
CO(g)		198
$CO_2(g)$	−394	213

(a) Calculate the standard enthalpy change of formation of carbon monoxide. [3]

ΔH^θ_f = ... kJ mol^{-1}

(b) Explain why the standard entropies of carbon dioxide and carbon monoxide are significantly greater than those of iron(III) oxide and iron. [1]

....

....

(c) The standard entropy change for this reaction, ΔS^θ, is +9 J K^{-1} mol^{-1}.

(i) Calculate the free energy change, ΔG^θ, for this reaction at 298K. [2]

ΔG^θ = ... kJ mol^{-1}

(ii) Explain why this reaction is feasible at all temperatures. [2]

(iii) Many industrial processes use high temperatures even when the reaction is feasible at low temperatures. Suggest why high temperatures are used. [1]

Total [9]

9

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END OF QUESTION PACK

5 questions · 63 marks · ~1 h 41 min
Source: WJEC CH5 (2008 modular spec, Jun 2010 – Jun 2016)
Curated for WJEC Chemistry 2015 spec A2 Unit 3 — Topic 5 (3.6 & 3.7)