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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Reactivity 1
What Drives Chemical Reactions?

Paper 2 Style Practice · Short-answer and extended-response questions

Medium to hard exam-style questions on Reactivity 1.1 – 1.4, with a full markscheme

Standard Level & Higher Level

Timing and marks

Higher Level: answer all five questions. Total **64 marks**. Suggested time **about 105 minutes**.

Standard Level: answer Questions 1–3 and omit Questions 4 and 5 (HL). Total **41 marks**. Suggested time **about 75 minutes**.

These timings use the same rate as the real examination (SL Paper 2: 50 marks in 1 h 30; HL Paper 2: 90 marks in 2 h 30). Reactivity 1 contains a large amount of additional higher level material, so the HL paper is longer.

Inside this booklet

1. Question 1 — Measuring enthalpy changes (12 marks)
2. Question 2 — Bond enthalpies and Hess's law (14 marks)
3. Question 3 — Energy from fuels (15 marks)
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5. Question 5 — HL: entropy, Gibbs energy and equilibrium (12 marks)
6. Markscheme and worked solutions

Instructions

- Work under examination conditions: set a timer, and do not look at the markscheme until you have finished.
- You need a **calculator** and a copy of the **IB Chemistry data booklet** (first assessment 2025). Where a question gives data, use those values. c(water) = 4.18 J g^{-1} K^{-1}; R = 8.31 J K^{-1} mol^{-1}.
- Write your answers in the spaces provided. Show all working in calculations and give answers with units and an appropriate number of significant figures.
- The number of marks for each part is shown in brackets. As a guide, one mark usually requires one distinct, correct point.

Question 1 Measuring enthalpy changes (12 marks)

(a) Distinguish between heat and temperature. **[2]**

A student added 0.120 g of magnesium ribbon to 50.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid in a polystyrene cup. The temperature of the solution rose by 10.4 K.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

(b) (i) Show that the hydrochloric acid is in excess. **[1]**

(ii) Calculate the enthalpy change of the reaction, in kJ mol^{-1}. **[3]**

(iii) The literature value is −467 kJ mol^{-1}. Suggest one reason why the student's value is less exothermic, and one specific improvement to the method that would reduce this error. **[2]**

(c) Sketch a labelled potential energy profile for this reaction. Show ΔH and the activation energy, E_a. **[3]**

(d) Deduce, with a reason, whether the reactants or the products are more energetically stable. **[1]**

Question 2 Bond enthalpies and Hess's law **(14 marks)**

Methanol can be made from carbon monoxide and hydrogen: $CO(g) + 2H_2(g) \rightarrow CH_3OH(g)$.

Bond	C≡O (in CO)	H–H	C–H	C–O	O–H
Average bond enthalpy / kJ mol^{-1}	1077	436	414	358	463

(a) Define the term *average bond enthalpy*. **[2]**

(b) Use the bond enthalpies to calculate the enthalpy change for the formation of gaseous methanol from CO and H_2. **[3]**

The enthalpy changes of the following reactions can be measured.

(1) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$	$\Delta H = -283$ kJ mol^{-1}
(2) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	$\Delta H = -286$ kJ mol^{-1}
(3) $CH_3OH(l) + 1\frac{1}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	$\Delta H = -726$ kJ mol^{-1}

(c) Use Hess's law to calculate ΔH for $CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$. **[3]**

(d) Suggest two reasons why your answers to (b) and (c) differ. **[2]**

(e) State Hess's law and explain why it is a consequence of the law of conservation of energy. **[2]**

(f) Explain why the carbon–oxygen bond in CO is much stronger and shorter than the carbon–oxygen bond in methanol. **[2]**

Question 3 Energy from fuels **(15 marks)**

Butane, C_4H_{10} (M = 58.14 g mol^{-1}, ΔH_c = −2878 kJ mol^{-1}), is used in camping stoves.

(a) (i) Write an equation for the complete combustion of butane. **[1]**

(ii) Write an equation for the incomplete combustion of butane to form carbon monoxide and water. **[1]**

(b) Explain why using a butane stove inside a closed tent is dangerous. **[2]**

(c) Calculate the specific energy of butane and the mass of CO_2 released per 1000 kJ of energy. Methane releases 49.4 g of CO_2 per 1000 kJ. Use your answer to comment on which fuel has the smaller effect on climate per unit of energy. **[3]**

(d) Explain how increasing concentrations of carbon dioxide in the atmosphere lead to global warming. **[3]**

Bioethanol is made by fermentation of glucose from sugar cane.

(e) (i) Write the equation for the fermentation of glucose. **[1]**

(ii) Evaluate the claim that bioethanol is a carbon-neutral fuel. **[2]**

(f) Methanol can be used in a fuel cell with an acidic electrolyte. Deduce the half-equations at the anode and at the cathode. **[2]**

HL Question 4 ΔH_f, ΔH_c and Born–Haber cycles (11 marks)

(a) Define the standard enthalpy change of formation. **[2]**

In some natural gas processing plants, hydrogen sulfide is removed by reaction with sulfur dioxide: $2H_2S(g) + SO_2(g) \rightarrow 3S(s) + 2H_2O(l)$. ΔH_f / kJ mol^{-1}: $H_2S(g)$ −20.6, $SO_2(g)$ −296.8, $H_2O(l)$ −285.8.

(b) (i) State why ΔH_f of S(s) is zero. **[1]**

(ii) Calculate ΔH for the reaction. **[2]**

(c) Calculate ΔH_f of propan-1-ol, $C_3H_7OH(l)$, using ΔH_c / kJ mol^{-1}: C(graphite) −394, $H_2(g)$ −286, $C_3H_7OH(l)$ −2021. **[2]**

Data for calcium chloride / kJ mol^{-1}: $\Delta H_f(CaCl_2)$ −796; $\Delta H_{atom}(Ca)$ +178; $IE_1(Ca)$ +590; $IE_2(Ca)$ +1145; E(Cl–Cl) +242; EA(Cl) −349.

(d) (i) Write an equation, with state symbols, for the second ionization energy of calcium. **[1]**

(ii) Calculate the lattice enthalpy of calcium chloride. **[2]**

(e) The lattice enthalpy of $MgCl_2$ is +2523 kJ mol^{-1}. Explain why it is larger than that of $CaCl_2$. **[1]**

HL Question 5 Entropy, Gibbs energy and equilibrium (12 marks)

Solid ammonium chloride decomposes on heating: $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$.

Substance	$NH_4Cl(s)$	$NH_3(g)$	$HCl(g)$
$\Delta H_f^\ominus$ / kJ mol^{-1}	−314.4	−46.1	−92.3
$S^\ominus$ / J K^{-1} mol^{-1}	94.6	192.5	186.9

(a) Predict, with a reason, the sign of the entropy change for the reaction. **[1]**

(b) Calculate $\Delta H^\ominus$ and $\Delta S^\ominus$ for the reaction. **[2]**

(c) Calculate $\Delta G^\ominus$ at 298 K and state whether the reaction is spontaneous. **[2]**

(d) Determine the temperature above which the decomposition is spontaneous. **[1]**

(e) Calculate the equilibrium constant, K, at 298 K, and comment on its value. **[2]**

(f) Calculate $\Delta G^\ominus$ and K at 700 K. **[2]**

(g) A mixture at 700 K has a reaction quotient Q = 100. Calculate ΔG for this mixture and deduce the direction in which the reaction will proceed. **[2]**

End of questions

Markscheme and worked solutions

Bold **[1]** shows where each mark is awarded. Accept equivalent wording that conveys the same chemistry. Text in italics gives examiner-style guidance. Allow follow-through (ECF) from an earlier error unless stated.

Question 1 (12 marks)

(a) Heat is energy transferred between two bodies because of a temperature difference (measured in J) **[1]**; temperature is a measure of the average kinetic energy of the particles (measured in K or °C) **[1]**.

(b) (i) n(Mg) = 0.120 / 24.31 = 4.94×10^{-3} mol, which needs 9.87×10^{-3} mol HCl; n(HCl) = $1.00 \times 0.0500 = 0.0500$ mol, so HCl is in excess **[1]**.

(ii) Q = $50.0 \times 4.18 \times 10.4 = 2174$ J **[1]**; n(Mg) = 4.94×10^{-3} mol (the limiting reactant) **[1]**; $\Delta H = -2.174 / 4.94 \times 10^{-3} =$ **-440 kJ mol^{-1}** (negative sign required) **[1]**.
Using the mass of magnesium in $Q = mc\Delta T$ scores no marks for that step.

(iii) Heat was lost to the surroundings (or absorbed by the cup and thermometer, or carried away by the escaping hydrogen gas and water vapour) **[1]**; for example: use a lid and insulate the cup, or record the temperature over time and extrapolate the cooling curve back to the time of adding the magnesium **[1]**.

(c) Axes labelled potential energy (y) and reaction coordinate (x), with reactants (Mg + 2HCl) above products ($MgCl_2 + H_2$) **[1]**; ΔH arrow pointing down from the reactant level to the product level **[1]**; single hump with E_a arrow from the reactant level to the top of the curve **[1]**.

(d) The products are more stable, because they have lower enthalpy than the reactants (the reaction is exothermic) **[1]**.

Question 2 (14 marks)

(a) The energy needed to break one mole of a bond in gaseous molecules (under standard conditions) **[1]**, averaged over a range of similar compounds **[1]**.

(b) bonds broken: C≡O + 2 H–H = 1077 + 2(436) = 1949 kJ **[1]**; bonds formed: 3 C–H + C–O + O–H = 3(414) + 358 + 463 = 2063 kJ **[1]**; $\Delta H = 1949 - 2063 =$ **-114 kJ mol^{-1}** **[1]**.

(c) Use (1) as written, (2) × 2, and reverse (3) **[1]**; $\Delta H = -283 + 2(-286) + 726$ **[1]** = **-129 kJ mol^{-1}** **[1]**.

(d) Bond enthalpies are average values, not the exact values for the bonds in CO and CH_3OH **[1]**; the answer to (c) forms liquid methanol, so it includes the energy released when gaseous methanol condenses (intermolecular forces), whereas bond enthalpies apply only to gases **[1]**.

(e) The enthalpy change for a reaction is independent of the route taken, provided the initial and final states are the same **[1]**. Energy cannot be created or destroyed, so if two routes had different enthalpy changes, going one way and returning the other would create energy **[1]**.

(f) CO has a triple bond (three shared pairs), whereas methanol has a C–O single bond **[1]**; more electron density between the nuclei gives a stronger attraction, pulling the nuclei closer together and requiring more energy to separate them **[1]**.

Question 3 (15 marks)

(a) (i) $C_4H_{10} + 6½O_2 \rightarrow 4CO_2 + 5H_2O$ (or $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$) **[1]**.
(ii) $C_4H_{10} + 4½O_2 \rightarrow 4CO + 5H_2O$ (or $2C_4H_{10} + 9O_2 \rightarrow 8CO + 10H_2O$) **[1]**.

(b) In a closed space the oxygen supply becomes limited, so incomplete combustion produces carbon monoxide **[1]**; CO is toxic because it binds to haemoglobin more strongly than oxygen, reducing the oxygen carried by the blood (it is also colourless and odourless, so it is not noticed) **[1]**.

(c) specific energy = 2878 / 58.14 = **49.5 kJ g^{-1} [1]**; 1000 kJ needs 1000 / 2878 = 0.347 mol butane, giving 4 × 0.347 = 1.39 mol CO_2 = **61.2 g [1]**; methane releases less CO_2 per kJ, because it has a lower percentage of carbon (higher H : C ratio), so it has the smaller effect on climate per unit of energy **[1]**.

(d) Short-wavelength radiation from the Sun passes through the atmosphere and warms the Earth's surface, which emits longer-wavelength infrared radiation **[1]**; CO_2 molecules absorb this IR radiation (their bonds vibrate, changing the dipole moment) **[1]**; the energy is re-radiated in all directions, partly back towards the Earth, so more CO_2 traps more heat and the average temperature rises **[1]**.

(e) (i) $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ **[1]**.
(ii) In favour: the CO_2 released by fermentation and combustion was absorbed by the sugar cane during photosynthesis only months earlier, so in principle there is no net increase **[1]**. Against: fossil fuels are used to make fertilizers, run farm machinery, transport the crop and distil the ethanol, so there is a net release of CO_2; the claim is not fully true **[1]**.

(f) Anode: $CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$ **[1]**; cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ (or $1½O_2 + 6H^+ + 6e^- \rightarrow 3H_2O$) **[1]**.

HL Question 4 (11 marks)

(a) The enthalpy change when one mole of a substance is formed from its elements **[1]** in their standard states under standard conditions (100 kPa, usually 298 K) **[1]**.

(b) (i) S(s) is an element in its standard state **[1]**.
(ii) products: 3(0) + 2(−285.8) = −571.6; reactants: 2(−20.6) + (−296.8) = −338.0 **[1]**; ΔH = −571.6 − (−338.0) = **−233.6 kJ [1]**.

(c) 3C(graphite) + $4H_2(g)$ + ½$O_2(g) \rightarrow C_3H_7OH(l)$; $\Delta H_f = \Sigma\Delta H_c$(reactants) − $\Sigma\Delta H_c$(products) = [3(−394) + 4(−286)] − (−2021) **[1]** = −1182 − 1144 + 2021 = **−305 kJ mol^{-1} [1]**.

(d) (i) $Ca^+(g) \rightarrow Ca^{2+}(g) + e^-$ **[1]**.
(ii) $\Delta H_{lat} = -\Delta H_f + \Delta H_{atom}(Ca) + IE_1 + IE_2 + E(Cl–Cl) + 2EA(Cl)$ = 796 + 178 + 590 + 1145 + 242 + 2(−349) **[1]** = **+2253 kJ mol^{-1} [1]**.
The full Cl–Cl bond enthalpy is used because two Cl atoms are needed; the electron affinity is doubled.

(e) Mg^{2+} has a smaller ionic radius than Ca^{2+} (same charge), so the ions are closer together and the electrostatic attraction to Cl^- is stronger **[1]**.

HL Question 5 (12 marks)

(a) Positive: a solid forms two moles of gas, which is far more dispersed **[1]**.

(b) $\Delta H^\ominus$ = (−46.1 − 92.3) − (−314.4) = **+176.0 kJ mol^{-1} [1]**; $\Delta S^\ominus$ = (192.5 + 186.9) − 94.6 = **+284.8 J K^{-1} mol^{-1} [1]**.

(c) $\Delta G^\ominus$ = 176.0 − 298 × 0.2848 **[1]** = **+91.1 kJ mol^{-1}**; positive, so not spontaneous at 298 K **[1]**.
Award [0] for the calculation if ΔS is used in J without conversion.

(d) T = $\Delta H/\Delta S$ = 176.0 / 0.2848 = **618 K** (345 °C) **[1]**.

(e) ln K = −91 100 / (8.31 × 298) = −36.8 **[1]**; K = **1.0 × 10^{-16}**: the equilibrium lies almost entirely on the side of the reactant, so ammonium chloride does not decompose at room temperature **[1]**.

(f) $\Delta G^\ominus$ = 176.0 − 700 × 0.2848 = **−23.4 kJ mol^{-1} [1]**; K = $e^{23\,400/(8.31 \times 700)} = e^{4.02}$ = **56** (accept 55–56) **[1]**.

(g) $\Delta G = \Delta G^{\ominus} + RT \ln Q = -23\,400 + 8.31 \times 700 \times \ln 100 = -23\,400 + 26\,790 =$ **+3.4 kJ mol^{-1} [1]**; ΔG is positive ($Q > K$), so the reverse reaction is spontaneous: NH_3 and HCl combine to form NH_4Cl until $Q = K$ **[1]**.

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