Chemistry SL

IB SL — Question Bank

Q1.

Worked Example 1 — Calculating A_r from isotope data

Chlorine has two naturally occurring isotopes:

- ^{35}Cl: abundance 75.77%
- ^{37}Cl: abundance 24.23%

$$A_r(\text{Cl}) = (0.7577 \times 35) + (0.2423 \times 37)$$
$$A_r(\text{Cl}) = 26.52 + 8.97 = 35.49 \approx 35.5$$

This matches the value you see on the periodic table. Note that A_r is never a whole number when two or more isotopes are present in significant amounts.

Common mistake: Using percentage instead of fractional abundance. Always divide percentages by 100 before substituting. Using 75.77 instead of 0.7577 gives an answer 100× too large.

Q2.

Worked Example 2 — Writing electron configurations

Write the full and condensed electron configurations for phosphorus ($Z = 15$).

Step 1: Total electrons = 15 (neutral atom).

Step 2: Fill in order: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Check: $2 + 2 + 6 + 2 + 3 = 15$ ✓

Step 3: Condensed — the noble gas before P is neon ($Z = 10$, configuration $1s^2\ 2s^2\ 2p^6$):

$[\text{Ne}]\ 3s^2\ 3p^3$

Applying Hund's rule to the $3p^3$: The three $3p$ electrons each occupy a separate orbital with parallel spins before any pairing occurs:

$3p$:↑ ↑ ↑

Q3.

Worked Example 3 — Mole calculations

Part (a): How many moles are in 9.80 g of sulfuric acid, $\mathrm{H_2SO_4}$?

Step 1 — Find molar mass: $M(\mathrm{H_2SO_4}) = 2(1.01) + 32.07 + 4(16.00) = 2.02 + 32.07 + 64.00 = 98.09\,\mathrm{g\,mol^{-1}}$

Step 2 — Apply $n = m/M$: $n = \frac{9.80}{98.09} = 0.0999 \approx 0.100\,\mathrm{mol}$

Part (b): What mass of sodium chloride (NaCl) contains 1.20×10^{24} formula units?

Step 1 — Find moles: $n = \frac{1.20 \times 10^{24}}{6.022 \times 10^{23}} = 1.99 \approx 2.00\,\mathrm{mol}$

Step 2 — Find mass: $m = n \times M = 2.00 \times 58.44 = 116.9 \approx 117\,\mathrm{g}$

Q4.

Worked Example 4 — Empirical and molecular formula from percentage composition

A compound has the following percentage composition by mass: C 40.00%, H 6.72%, O 53.28%. Its molar mass is $M_r = 90.08\,\mathrm{g\,mol^{-1}}$. Find its empirical and molecular formulae.

Step 1: Assume 100 g of compound — percentages become masses in grams:

- $m(\mathrm{C}) = 40.00\,\mathrm{g}, m(\mathrm{H}) = 6.72\,\mathrm{g}, m(\mathrm{O}) = 53.28\,\mathrm{g}$

Step 2: Convert to moles: $n(\mathrm{C}) = \frac{40.00}{12.01} = 3.331 \quad n(\mathrm{H}) = \frac{6.72}{1.008} = 6.667 \quad n(\mathrm{O}) = \frac{53.28}{16.00} = 3.330$

Step 3: Divide by the smallest value (3.330): $\mathrm{C}: \frac{3.331}{3.330} \approx 1 \quad \mathrm{H}: \frac{6.667}{3.330} \approx 2 \quad \mathrm{O}: \frac{3.330}{3.330} = 1$

Empirical formula: $\mathrm{CH_2O}$ — $M_r(\text{empirical}) = 12 + 2 + 16 = 30\,\mathrm{g\,mol^{-1}}$

Step 4: Find the multiple: $\text{multiple} = \frac{90.08}{30} = 3.003 \approx 3$

Molecular formula: $\mathrm{C_3H_6O_3}$

Q5.

Worked Example 5 — Ideal gas calculation

Calculate the volume occupied by 3.20 g of oxygen gas ($\mathrm{O_2}$) at 298 K and 101 kPa. Give your answer in litres.

Step 1: Find moles: $M(\mathrm{O_2}) = 2 \times 16.00 = 32.00\,\mathrm{g\,mol^{-1}}$

$n = \frac{3.20}{32.00} = 0.100\,\mathrm{mol}$

Step 2: Apply $PV = nRT$. Use $R = 8.314\,\text{J mol}^{-1}\text{K}^{-1}$, P in Pa:

$V = \frac{nRT}{P} = \frac{0.100 \times 8.314 \times 298}{101\,000} = \frac{247.8}{101\,000} = 2.45 \times 10^{-3}\,\text{m}^3$

Step 3: Convert to litres ($1\,\text{m}^3 = 1000\,\text{L}$):

$V = 2.45 \times 10^{-3} \times 1000 = 2.45\,\text{L}$

Q6.

Worked Example: Drawing a dot-and-cross diagram for an ionic compound

Draw the formation of MgCl_2 showing electron transfer.

Step 1: Identify the atoms and their valence electrons.

- Mg (Group 2): 2 valence electrons
- Cl (Group 17): 7 valence electrons. Two Cl atoms are needed.

Step 2: Mg transfers 1 electron to each Cl atom.

- Mg loses 2 electrons → forms Mg^{2+} (8 electrons remaining, neon configuration)
- Each Cl gains 1 electron → forms Cl^- (8 electrons in outer shell, argon configuration)

Step 3: Write the result.

$$\text{Mg} + 2\text{Cl} \rightarrow \text{Mg}^{2+} + 2\text{Cl}^-$$

Charge check: Overall charge = (2+) + 2(1−) = 0. The compound is electrically neutral. ✓

Q7.

Worked Example: Drawing the Lewis structure of H_2O

Step 1: Count valence electrons.

- O (Group 16): 6 electrons
- H × 2 (Group 1): 1 × 2 = 2 electrons
- **Total: 8 electrons**

Step 2: Arrange atoms. O is the central atom; both H atoms are terminal.

Step 3: Place single bonds: O–H and O–H. This uses 2 × 2 = 4 electrons. Remaining: 8 − 4 = 4 electrons.

Step 4: Complete octets on outer atoms. H already has 2 electrons (full duet). Place remaining 4 electrons on O as 2 lone pairs.

Step 5: Check. O has: 2 bonding pairs (4 electrons) + 2 lone pairs (4 electrons) = 8 electrons = full octet. ✓

Result: Water has 2 O–H single bonds and 2 lone pairs on oxygen.

Q8.

Worked Example: Drawing the Lewis structure of CO_2

Step 1: Count valence electrons.

- C (Group 14): 4 electrons
- O × 2 (Group 16): 6 × 2 = 12 electrons
- **Total: 16 electrons**

Step 2: C is the central atom; both O atoms are terminal (O=C=O).

Step 3: Place single bonds: C–O and C–O. Uses 4 electrons. Remaining: 12 electrons.

Step 4: Complete octets on O atoms first. Each O needs 6 more electrons = 3 lone pairs each. That would use 12 electrons total. But C only has 4 electrons (2 from each single bond) = not a full octet.

Step 5: C needs 4 more electrons to complete its octet. Convert 1 lone pair from each O into a bonding pair with C → forms two C=O double bonds.

Final check: C has 4 + 4 = 8 electrons (two double bonds). Each O has 1 double bond (4 electrons) + 2 lone pairs (4 electrons) = 8 electrons. Total electrons used: 4 (bonds) + 4 (bonds) + 4 (lone pairs on each O × 2) = 16. ✓

Result: CO_2 has two C=O double bonds and 2 lone pairs on each oxygen.

Q9.

Worked Example: Drawing the Lewis structure of NH_3

Step 1: Count valence electrons.

- N (Group 15): 5 electrons
- H × 3 (Group 1): 1 × 3 = 3 electrons
- **Total: 8 electrons**

Step 2: N is the central atom; all H atoms are terminal.

Step 3: Place single bonds: N–H, N–H, N–H. Uses 6 electrons. Remaining: 2 electrons.

Step 4: Complete octets on H (already done — each has 2). Place remaining 2 electrons on N as 1 lone pair.

Step 5: N has: 3 bonding pairs (6e) + 1 lone pair (2e) = 8 electrons = full octet. ✓

Result: Ammonia has 3 N–H single bonds and 1 lone pair on nitrogen.

Q10.

Worked Example: Drawing the Lewis structure of PCl_3

Step 1: Count valence electrons.

- P (Group 15): 5 electrons
- Cl × 3 (Group 17): 7 × 3 = 21 electrons
- **Total: 26 electrons**

Step 2: P is the central atom; all Cl atoms are terminal.

Step 3: Place 3 P–Cl single bonds. Uses 6 electrons. Remaining: 20 electrons.

Step 4: Complete octets on each Cl: each needs 6 more electrons = 3 lone pairs. 3 × 6 = 18 electrons used. Remaining: 2 electrons.

Step 5: Place remaining 2 electrons on P as 1 lone pair.

Check: P has 3 bonding pairs + 1 lone pair = 8 electrons = full octet. Each Cl has 3 lone pairs + 1 bonding pair = 8 electrons. Total: 6 + 18 + 2 = 26. ✓

Result: PCl_3 has 3 P–Cl single bonds, 1 lone pair on P, and 3 lone pairs on each Cl.

Q11.

Worked Example: Determining the shape of H_2O

Step 1: Draw the Lewis structure. O is central, 2 O–H bonds, 2 lone pairs on O.

Step 2: Count electron domains around O.

- 2 bonding pairs (the two O–H bonds)
- 2 lone pairs
- **Total: 4 electron domains**

Step 3: Electron geometry = tetrahedral (4 domains).

Step 4: Determine molecular shape. 2 lone pairs are present — they occupy two corners of the tetrahedron but are not "seen" in the shape name. The shape described by the 3 atoms (O and 2 H atoms) is **bent** (V-shaped).

Step 5: Bond angle. Two lone pairs compress the H–O–H angle below 109.5°. Bond angle ≈ 104.5°.

Q12.

Worked Example: Determining the shape of NH_3

Step 1: Lewis structure: N is central, 3 N–H single bonds, 1 lone pair on N.

Step 2: Electron domains around N: 3 bonding pairs + 1 lone pair = **4 domains**.

Step 3: Electron geometry = tetrahedral.

Step 4: Molecular shape: 1 lone pair is invisible in shape name. The 4 atoms (N + 3 H) form a **trigonal pyramidal** shape — like a triangular base with N at the apex.

Step 5: Bond angle: 1 lone pair pushes bonding pairs slightly → H–N–H angle ≈ 107°.

Q13.

Worked Example: Predicting the polarity of CCl_4

Step 1: Identify the bonds. C–Cl: carbon (EN ≈ 2.5) and chlorine (EN ≈ 3.0). The difference is ~0.5. Each C–Cl bond is polar ($C^{\delta+}$–$Cl^{\delta-}$).

Step 2: Determine the shape. C is the central atom, 4 Cl atoms, 0 lone pairs → **tetrahedral**, bond angles 109.5°.

Step 3: Assess symmetry. The four C–Cl dipoles point symmetrically from the centre to the four corners of the tetrahedron. By vector addition, the four equal and symmetrically arranged dipoles cancel exactly.

Q14.

Worked Example: Why does HF have a higher boiling point than HCl?

Step 1: Identify the IMFs in each molecule.

- HF: H is directly bonded to F → **hydrogen bonds** form between HF molecules (in addition to London forces and dipole–dipole forces).
- HCl: H is bonded to Cl. Cl is not N, O, or F → no hydrogen bonds. HCl is polar → **dipole–dipole forces** + London forces.

Step 2: Compare IMF strength.

- Hydrogen bonds (in HF) are significantly stronger than dipole–dipole forces (in HCl).
- HF has a smaller molar mass (20 g/mol) than HCl (36.5 g/mol), so HCl has stronger London forces — but this is outweighed by the hydrogen bonding in HF.

Step 3: Conclusion.

- More energy is needed to overcome the hydrogen bonds in HF than the dipole–dipole forces in HCl.
- Therefore HF has a higher boiling point than HCl despite having a lower molar mass.

Q15.

Question 1 [VSEPR and Polarity] [~7 marks]

Consider the molecules BF_3, NF_3, and CF_4.

(a) Draw the Lewis (electron dot) structure for each molecule, showing all bonding pairs and lone pairs.

(b) Use VSEPR theory to predict the shape and bond angle of each molecule.

(c) Explain why NF_3 is polar but BF_3 and CF_4 are non-polar, despite all three containing polar X-F bonds.

▶ Show Solution

Q16.

Question 2 [Intermolecular Forces and Properties] [~7 marks]

The boiling points of the hydrogen halides are:

Compound	Boiling point (K)
HF	293
HCl	188
HBr	206
HI	238

(a) Explain the general trend in boiling points from HCl to HI.

(b) Explain why HF has an anomalously high boiling point compared to the other hydrogen halides.

(c) Explain why NaCl (melting point 1074 K) has a much higher melting point than any of the hydrogen halides.

▶ Show Solution

Q17.

Question 3 [Bonding Types Comparison] [~6 marks]

Compare and contrast the bonding and structure in diamond, graphite, and ice (H_2O), and explain how the bonding in each substance accounts for its electrical conductivity.

▶ Show Solution

Q18.

Example: 50.0 cm^3 of 1.00 mol dm^{-3} HCl is mixed with 50.0 cm^3 of 1.00 mol dm^{-3} NaOH. The temperature rises from 22.0°C to 28.8°C. Calculate the enthalpy of neutralisation per mole.

Step 1: Total volume = 100.0 cm^3. Assume density = 1.00 g cm^{-3}, so $m = 100.0$ g.

Step 2: $\Delta T = 28.8 - 22.0 = 6.8$ K

Step 3: $q = mc\Delta T = 100.0 \times 4.18 \times 6.8 = 2842$ J $= 2.84$ kJ

Step 4: Moles of HCl = $0.0500 \times 1.00 = 0.0500$ mol

Step 5: $\Delta H = -\dfrac{q}{n} = -\dfrac{2.84}{0.0500} = -56.8 \text{ kJ mol}^{-1}$

The negative sign indicates the reaction is exothermic (temperature rose).

Q19.

Example: Calculate ΔH° for the reaction:

$$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$$

Given: $\Delta H_f^\circ[\text{CH}_4(\text{g})] = -74.8$, $\Delta H_f^\circ[\text{CO}_2(\text{g})] = -393.5$, $\Delta H_f^\circ[\text{H}_2\text{O}(\text{l})] = -285.8$ (all in kJ mol^{-1}).

Solution:

$$\Delta H^\circ_{\text{rxn}} = \left[\Delta H_f^\circ(\text{CO}_2) + 2 \times \Delta H_f^\circ(\text{H}_2\text{O})\right] - \left[\Delta H_f^\circ(\text{CH}_4) + 2 \times \Delta H_f^\circ(\text{O}_2)\right]$$

$$= \left[-393.5 + 2(-285.8)\right] - \left[-74.8 + 0\right]$$

$$= (-393.5 - 571.6) - (-74.8) = -965.1 + 74.8 = -890.3 \text{ kJ mol}^{-1}$$

The large negative value confirms this is a highly exothermic combustion reaction.

Q20.

Example: Calculate ΔH° for the reaction:

$$\text{C}(\text{s}) + 2\text{H}_2(\text{g}) \rightarrow \text{CH}_4(\text{g})$$

Given: $\Delta H_c^\circ[\text{C}(\text{s})] = -393.5$, $\Delta H_c^\circ[\text{H}_2(\text{g})] = -285.8$, $\Delta H_c^\circ[\text{CH}_4(\text{g})] = -890.3$ (all in kJ mol^{-1}).

Solution:

$$\Delta H^\circ_{\text{rxn}} = \left[\Delta H_c^\circ(\text{C}) + 2 \times \Delta H_c^\circ(\text{H}_2)\right] - \left[\Delta H_c^\circ(\text{CH}_4)\right]$$

$$= \left[-393.5 + 2(-285.8)\right] - \left[-890.3\right]$$

$$= -965.1 + 890.3 = -74.8 \text{ kJ mol}^{-1}$$

This equals the standard enthalpy of formation of methane, as expected from the equation.

Q21.

Example: Use bond enthalpies to estimate ΔH for the combustion of methane:

$$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$$

Bond enthalpies: C-H = 414, O=O = 498, C=O (in CO_2) = 804, O-H = 463 (all in kJ mol^{-1}).

Bonds broken (reactants):

- $4 \times$ C-H = $4 \times 414 = 1656$
- $2 \times$ O=O = $2 \times 498 = 996$
- **Total broken = 2652 kJ**

Bonds formed (products):

- $2 \times$ C=O = $2 \times 804 = 1608$
- $4 \times$ O-H = $4 \times 463 = 1852$
- **Total formed = 3460 kJ**

$\Delta H \approx 2652 - 3460 = -808 \text{ kJ mol}^{-1}$

The value is approximate because bond enthalpies are averages. The data booklet value (-890 kJ mol^{-1}) is more accurate.

Q22.

Example: Calculate ΔS° for the decomposition of calcium carbonate:

$\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

Given: $S^\circ[\text{CaCO}_3(\text{s})] = 92.9$, $S^\circ[\text{CaO}(\text{s})] = 39.7$, $S^\circ[\text{CO}_2(\text{g})] = 213.7$ (all in J K^{-1} mol^{-1}).

Solution:

$\Delta S^\circ = [39.7 + 213.7] - [92.9] = 253.4 - 92.9 = +160.5 \text{ J K}^{-1}\text{mol}^{-1}$

The large positive value makes sense: a gas (CO_2) is produced from a solid, greatly increasing disorder.

Q23.

Example: For the decomposition of calcium carbonate:

$\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

$\Delta H^\circ = +178 \text{ kJ mol}^{-1}$, $\Delta S^\circ = +160.5 \text{ J K}^{-1}\text{mol}^{-1}$

(a) Is this reaction spontaneous at 298 K?

$\Delta G = 178 - (298 \times 0.1605) = 178 - 47.8 = +130.2 \text{ kJ mol}^{-1}$

$\Delta G > 0$, so the reaction is **not spontaneous** at 298 K.

(b) At what temperature does it become spontaneous?

$T = \frac{\Delta H}{\Delta S} = \frac{178}{0.1605} = 1109 \text{ K} \approx 836^\circ\text{C}$

Above 1109 K (836°C), the entropy term ($T\Delta S$) becomes large enough to overcome the positive ΔH, making ΔG negative.

Q24.

Example: The Haber process:

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

$\Delta H^\circ = -92.2\ \mathrm{kJ\ mol^{-1}}$, $\Delta S^\circ = -198.7\ \mathrm{J\ K^{-1} mol^{-1}}$

This is an exothermic reaction ($\Delta H < 0$) with a decrease in entropy ($\Delta S < 0$, because 4 moles of gas become 2 moles). At low temperatures, $\Delta G < 0$ (spontaneous). At high temperatures, the $-T\Delta S$ term becomes a large positive number that overcomes the negative ΔH.

$T = \frac{-92.2}{-0.1987} = 464\ \mathrm{K} \approx 191^\circ\mathrm{C}$

Above 464 K, $\Delta G > 0$ and the forward reaction is non-spontaneous. This is why the Haber process uses a compromise temperature (around 450°C) — high enough for a reasonable rate but near the thermodynamic limit.

Q25.

Worked Example 1.1 — Reading the Maxwell-Boltzmann Distribution

A reaction has an activation energy of 50 kJ/mol. At 25 °C, 8% of molecules have energy ≥ E_a. At 35 °C, 16% of molecules have energy ≥ E_a.

Q: By what factor has the rate approximately increased?

Solution:

The fraction of molecules exceeding E_a has doubled (from 8% to 16%). Since all other factors (concentration, surface area) are unchanged, the rate of effective collisions approximately doubles.

Rate factor ≈ 2

This illustrates the general rule of thumb: a 10 °C rise roughly doubles the rate for many reactions — though this depends on E_a and is not a precise universal law.

Q26.

Practice MCQ — Maxwell-Boltzmann Interpretation

The diagram below shows a Maxwell-Boltzmann distribution for a gas at temperature T_1, with the activation energy E_a marked. The same sample is then heated to a higher temperature T_2. Which statement correctly describes the curve at T_2 compared with T_1?

(A) The peak is higher and shifted to lower energy; the area beyond E_a decreases
(B) The peak is lower and shifted to higher energy; the area beyond E_a increases

(C) The peak is unchanged; the position of E_a moves to the left **(D)** The total area under the curve increases because more molecules are present

Answer: (B) Raising the temperature flattens and broadens the distribution: the peak drops and shifts to higher energy, and the tail extends further right, so a greater proportion of molecules have energy $\geq E_a$ (larger shaded area beyond E_a). The total area under the curve stays the same because the number of molecules is unchanged (D is wrong), and E_a itself does not move — only a catalyst would lower it (C is wrong).

Q27.

Worked Example 2.1 — Calculating Average Rate

The concentration of a reactant A was measured over time:

Time / s	$[\mathrm{A}]$ / mol dm^{-3}
0	0.80
20	0.60
40	0.44
60	0.32
80	0.24

Q(a): Calculate the average rate of reaction between $t = 0$ and $t = 40$ s.

Q(b): Calculate the average rate between $t = 40$ s and $t = 80$ s.

Q(c): What does the change in rate tell you about the reaction as it proceeds?

Solution (a):

$$\text{average rate} = \frac{\Delta[\mathrm{A}]}{\Delta t} = \frac{0.80 - 0.44}{40 - 0} = \frac{0.36}{40} = 9.0 \times 10^{-3} \text{ mol dm}^{-3} \text{ s}^{-1}$$

Solution (b):

$$\text{average rate} = \frac{0.44 - 0.24}{80 - 40} = \frac{0.20}{40} = 5.0 \times 10^{-3} \text{ mol dm}^{-3} \text{ s}^{-1}$$

Solution (c): The rate has decreased (from 9.0×10^{-3} to 5.0×10^{-3} mol dm^{-3} s^{-1}). This is because as the reaction proceeds, the concentration of A decreases. Fewer particles per unit volume means fewer collisions per second, so the rate falls. This is why concentration-time graphs are curves (decreasing gradient) rather than straight lines.

Q28.

Worked Example 3.1 — Catalyst Explanation

Q: A reaction $\mathrm{A} \rightarrow \mathrm{B}$ has $E_a = 120$ kJ mol^{-1}. A catalyst reduces this to 80 kJ mol^{-1}.

Explain, with reference to the Maxwell-Boltzmann distribution, why the rate increases.

Q29.

Worked Example 4.1 — Reading Rate from a Graph

A reaction produces gas. The volume of gas collected is measured over time:

Time / s	Volume of gas / cm^3
0	0
10	18
20	32
30	42
40	48
50	51
60	52

Q(a): Calculate the average rate of gas production between $t = 0$ and $t = 20$ s.

Q(b): Between $t = 40$ s and $t = 60$ s, the rate is much slower. Suggest why.

Q(c): At what point does the reaction appear to stop? Explain.

Solution (a):

$$\text{average rate} = \frac{32-0}{20-0} = \frac{32}{20} = 1.6 \text{ cm}^3 \text{ s}^{-1}$$

Solution (b): Between $t = 40$ and $t = 60$ s, only 4 cm^3 more gas is produced (compared to 32 cm^3 in the first 20 s). By this point, the reactants have been largely consumed — much lower concentration means fewer collisions per second, so the rate has fallen dramatically.

Solution (c): The reaction appears to stop at approximately $t = 55$–60 s, where the volume becomes constant at ~52 cm^3. This means no more gas is being produced — either a reactant has been fully consumed (limiting reagent exhausted) or, if this is a reversible reaction, equilibrium has been reached.

Q30.

Worked Example 5.1 — Le Chatelier's Principle

Consider the equilibrium:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \quad \Delta H = -197 \text{ kJ mol}^{-1}$$

Predict the direction of shift for each change:

(a) Temperature is increased from 500 °C to 600 °C.

(b) Pressure is doubled.

(c) Some SO_3 is removed from the equilibrium mixture.

(d) A vanadium(V) oxide (V_2O_5) catalyst is added.

Solutions:

(a) The forward reaction is exothermic ($\Delta H = -197$ kJ mol^{-1}). Increasing temperature is equivalent to adding heat (a product in the exothermic direction). The system opposes this by shifting **left** (reverse), decreasing $[SO_3]$ and increasing $[SO_2]$ and $[O_2]$. K_c decreases.

(b) Left side: 2 + 1 = **3 moles of gas**. Right side: **2 moles of gas**. Increasing pressure favours the side with fewer moles of gas. Shift **right** (forward), producing more SO_3. K_c unchanged.

(c) Removing product SO_3 reduces $[SO_3]$. System shifts **right** (forward) to replace the removed product, consuming SO_2 and O_2. K_c unchanged.

(d) The catalyst (V_2O_5) speeds up both forward and reverse reactions equally. **No shift.** Equilibrium is reached faster, but the equilibrium position and K_c are unchanged.

Q31.

Practice MCQ — Le Chatelier Prediction

Consider the endothermic equilibrium:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq) \quad \Delta H > 0$$

Which change will shift the equilibrium to the **right** (producing more H_3O^+)?

(A) Decreasing the temperature **(B)** Adding a few drops of concentrated CH_3COOH **(C)** Adding sodium ethanoate, which increases $[CH_3COO^-]$ **(D)** Adding a catalyst

Answer: (B) Adding more CH_3COOH raises the concentration of a reactant, so the system shifts right to oppose the increase, producing more H_3O^+. (A) is wrong because lowering the temperature of an endothermic forward reaction shifts it left. (C) raises a product concentration, shifting left. (D) a catalyst speeds both directions equally and never shifts the position. Note that H_2O is a pure liquid and does not appear in the equilibrium reasoning.

Q32.

Practice MCQ — Writing the K_c Expression

Which expression correctly gives K_c for the heterogeneous equilibrium below?

$$2C(s) + O_2(g) \rightleftharpoons 2CO(g)$$

(A) $K_c = \dfrac{[CO]^2}{[C]^2[O_2]}$ **(B)** $K_c = \dfrac{[CO]^2}{[O_2]}$ **(C)** $K_c = \dfrac{[O_2]}{[CO]^2}$ **(D)** $K_c = \dfrac{2[CO]}{[O_2]}$

Answer: (B) Pure solids are excluded from the K_c expression because their concentration is effectively constant, so carbon does not appear (eliminating A). Products go on top and reactants on the bottom (eliminating C, which is inverted), and

the coefficient 2 of CO becomes the exponent rather than a multiplier (eliminating D). This leaves $K_c = \dfrac{[\mathrm{CO}]^2}{[\mathrm{O_2}]}$.

Q33.

Worked Example 6.1 — K_c from Equilibrium Concentrations

At 500 K, the following equilibrium is established:

$\mathrm{H_2(g) + I_2(g) \rightleftharpoons 2HI(g)}$

At equilibrium: $[\mathrm{H_2}] = 0.10$ mol dm^{-3}, $[\mathrm{I_2}] = 0.10$ mol dm^{-3}, $[\mathrm{HI}] = 0.74$ mol dm^{-3}

Calculate K_c.

Solution:

Write the K_c expression:

$K_c = \frac{[\mathrm{HI}]^2}{[\mathrm{H_2}][\mathrm{I_2}]}$

Substitute equilibrium concentrations:

$K_c = \frac{(0.74)^2}{(0.10)(0.10)} = \frac{0.5476}{0.0100} = 54.8$

$K_c = 54.8$ (no units for this equilibrium since the powers cancel: $\frac{\text{mol}^2\ \text{dm}^{-6}}{\text{mol}^2\ \text{dm}^{-6}}$)

Interpretation: $K_c = 54.8 \gg 1$, so products (HI) are favoured at this temperature. Most of the hydrogen and iodine has been converted to hydrogen iodide at equilibrium.

Q34.

Worked Example 6.2 — ICE Table to Find Equilibrium Concentration

At a certain temperature, $K_c = 4.0$ for the reaction:

$\mathrm{A(g) + B(g) \rightleftharpoons C(g) + D(g)}$

0.80 mol of A and 0.80 mol of B are placed in a 1.00 dm^3 flask. Calculate the equilibrium concentrations of all species.

Solution — ICE Table:

The ICE table tracks Initial concentrations, Change, and Equilibrium concentrations:

Species	[A]	[B]	[C]	[D]
Initial / mol dm^{-3}	0.80	0.80	0	0
Change / mol dm^{-3}	$-x$	$-x$	$+x$	$+x$
Equilibrium / mol dm^{-3}	$0.80 - x$	$0.80 - x$	x	x

Write the K_c expression and substitute:

$K_c = \frac{[\text{C}][\text{D}]}{[\text{A}][\text{B}]} = \frac{(x)(x)}{(0.80-x)(0.80-x)} = \frac{x^2}{(0.80-x)^2} = 4.0$

Take the square root of both sides:

$\frac{x}{0.80-x} = \sqrt{4.0} = 2.0$

Solve for x:

$x = 2.0(0.80 - x) = 1.60 - 2.0x$

$3.0x = 1.60$

$x = 0.533 \text{ mol dm}^{-3}$

Equilibrium concentrations:

- $[\text{A}] = [\text{B}] = 0.80 - 0.533 = \mathbf{0.27}$ mol dm^{-3}
- $[\text{C}] = [\text{D}] = x = \mathbf{0.53}$ mol dm^{-3}

Check: $K_c = \frac{(0.533)^2}{(0.267)^2} = \frac{0.284}{0.0713} \approx 4.0$ ✓

Q35.

Worked Example 7.1 — Haber Process K_c Calculation

At 500 °C, the equilibrium concentrations in a Haber Process reactor are:

$[\text{N}_2] = 3.0$ mol dm^{-3}, $[\text{H}_2] = 9.0$ mol dm^{-3}, $[\text{NH}_3] = 2.4$ mol dm^{-3}

Calculate K_c at this temperature.

Solution:

$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$

$K_c = \frac{(2.4)^2}{(3.0)(9.0)^3} = \frac{5.76}{(3.0)(729)} = \frac{5.76}{2187} = 2.63 \times 10^{-3}$

$K_c = 2.63 \times 10^{-3}$ at 500 °C

Interpretation: $K_c \ll 1$, confirming that at 500 °C, the equilibrium lies to the left — at this temperature, only a small fraction of N_2 and H_2 is converted to NH_3 at equilibrium (~15% yield). This is why the Haber Process recycles unreacted gases.

Q36.

Q1. A catalyst is added to a reaction mixture at constant temperature. Which statement correctly describes the effect of the catalyst on the Maxwell-Boltzmann distribution and activation energy?

(A) The distribution shifts to higher energies and E_a decreases **(B)** The distribution remains unchanged and E_a decreases **(C)** The distribution shifts to higher energies and E_a remains unchanged **(D)** Both the distribution and E_a remain unchanged

Answer: (B) A catalyst lowers the activation energy by providing an alternative pathway, but does not change the temperature — the Maxwell-Boltzmann distribution is determined only by temperature and remains unchanged.

Q37.

Q2. For the equilibrium $\mathrm{N_2O_4(g)} \rightleftharpoons 2\mathrm{NO_2(g)}$, the equilibrium constant expression is:

(A) $K_c = \dfrac{[\mathrm{NO_2}]}{[\mathrm{N_2O_4}]}$ **(B)** $K_c = \dfrac{[\mathrm{NO_2}]^2}{[\mathrm{N_2O_4}]}$ **(C)** $K_c = \dfrac{2[\mathrm{NO_2}]}{[\mathrm{N_2O_4}]}$ **(D)** $K_c = \dfrac{[\mathrm{N_2O_4}]}{[\mathrm{NO_2}]^2}$

Answer: (B) Products over reactants, each raised to its stoichiometric coefficient: $K_c = \dfrac{[\mathrm{NO_2}]^2}{[\mathrm{N_2O_4}]}$. The coefficient 2 becomes an exponent, not a multiplier.

Q38.

Question 1 [Collision Theory and Rate Factors] [~6 marks]

The reaction between hydrochloric acid and magnesium ribbon is studied:

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

(a) Using collision theory, explain why increasing the concentration of HCl increases the rate of this reaction.

(b) Explain why increasing the temperature increases the rate of reaction. Your answer should refer to the Maxwell-Boltzmann distribution.

(c) A catalyst is added. Explain how a catalyst increases the rate without being consumed.

▶ Show Solution

Q39.

Question 2 [K_c Calculation] [~7 marks]

The equilibrium reaction below was studied at 450 °C:

$$\mathrm{H_2}(g) + \mathrm{I_2}(g) \rightleftharpoons 2\mathrm{HI}(g)$$

At equilibrium, the concentrations in a 1.0 $\mathrm{dm^3}$ container are: $[\mathrm{H_2}] = 0.11\ \mathrm{mol\ dm^{-3}}$, $[\mathrm{I_2}] = 0.11\ \mathrm{mol\ dm^{-3}}$, $[\mathrm{HI}] = 0.78\ \mathrm{mol\ dm^{-3}}$.

(a) Write the expression for K_c for this reaction.

(b) Calculate the value of K_c at 450 °C.

(c) If the temperature is increased and K_c decreases, deduce whether the forward reaction is exothermic or endothermic. Explain your reasoning.

▶ Show Solution

Q40.

Question 3 [Le Chatelier's Principle] [~7 marks]

The Haber process for the manufacture of ammonia is represented by:

$$\mathrm{N_2}(g) + 3\mathrm{H_2}(g) \rightleftharpoons 2\mathrm{NH_3}(g) \quad \Delta H = -92 \text{ kJ mol}^{-1}$$

Industrial conditions: 400-500 °C, ~200 atm, iron catalyst.

(a) Use Le Chatelier's principle to predict and explain the effect of increasing pressure on the yield of ammonia.

(b) The forward reaction is exothermic. According to Le Chatelier's principle, a low temperature would favour ammonia production. Explain why the industrial process uses 400-500 °C despite this.

(c) Explain why the iron catalyst does not affect the position of equilibrium or the value of K_c.

▶ Show Solution

Q41.

Worked Example 1 — Explaining reactivity trend in Group 1

Question: Explain why potassium reacts more vigorously with water than lithium does.

Q42.

Worked Example 2 — Identifying functional groups

The following four compounds have the molecular formula $\mathrm{C_3H_6O}$:

- Compound A: $\mathrm{CH_3CH_2CHO}$ (propanal)
- Compound B: $\mathrm{CH_3COCH_3}$ (propanone)
- Compound C: $\mathrm{CH_2{=}CHCH_2OH}$ (prop-2-en-1-ol)
- Compound D: $\mathrm{CH_3CH{=}O}$ [same as A — ethanal would be $\mathrm{C_2H_4O}$; this is incorrect — ignore]

Classify the functional group in A and B and explain how you would distinguish them experimentally.

Compound A — propanal: The -CHO group at the end of the chain identifies this as an **aldehyde**.

Compound B — propanone: The C=O group is between two methyl groups (internal position) — this is a **ketone**.

Experimental distinction: Add Tollens' reagent (ammoniacal $\mathrm{AgNO_3}$) to each compound and warm gently.

- Compound A (aldehyde) will be **oxidised** to propanoic acid ($\mathrm{CH_3CH_2COOH}$), and a **silver mirror** will form on the inside of the test tube.
- Compound B (ketone) will show **no reaction** with Tollens' reagent — ketones resist oxidation.

Q43.

Worked Example: Calculating pH

What is the pH of a $0.01\ \mathrm{mol/dm^3}$ solution of HCl?

Step 1: HCl is a strong acid, so it fully ionizes: $[\mathrm{H^+}] = 0.01 = 10^{-2}\ \mathrm{mol/dm^3}$

Step 2: $\mathrm{pH} = -\log(10^{-2}) = 2$

Q44.

Worked Example: Finding $[\mathrm{H^+}]$ from pH

A solution has pH 4.5. What is the hydrogen ion concentration?

$[\mathrm{H^+}] = 10^{-\mathrm{pH}} = 10^{-4.5} = 3.16 \times 10^{-5}\ \mathrm{mol/dm^3}$

Q45.

Q45. Which of the following is a conjugate acid-base pair?

A. HCl and NaOH

B. $\mathrm{H_2SO_4}$ and $\mathrm{SO_4^{2-}}$

C. $\mathrm{NH_3}$ and $\mathrm{NH_4^+}$

D. HCl and $\mathrm{HNO_3}$

Q46.

Q46. What is the pH of a $0.001\ \mathrm{mol/dm^3}$ solution of $\mathrm{HNO_3}$?

A. 1

B. 2

C. 3

D. 4

Q47.

Q47. At the equivalence point of a titration between ethanoic acid and sodium hydroxide, the pH is:

A. Exactly 7

B. Below 7

C. Above 7

D. Impossible to determine

Q48.

Q48. Which statement correctly distinguishes a strong acid from a weak acid?

A. A strong acid has a higher concentration than a weak acid

B. A strong acid reacts faster than a weak acid

C. A strong acid ionizes completely in water; a weak acid ionizes partially

D. A strong acid always has a lower pH than a weak acid

Q49.

Q49. A student adds 25.0 cm^3 of $0.100\ mol/dm^3$ $NaOH$ to 25.0 cm^3 of $0.100\ mol/dm^3$ CH_3COOH. Which best describes the pH at the equivalence point?

A. pH = 7, because equal moles of acid and base have reacted

B. pH < 7, because CH_3COOH is a stronger acid than $NaOH$ is a base

C. pH > 7, because the salt CH_3COONa produces a basic solution

D. pH = 7, because the solution is neutral after neutralization

Q50.

Q50. Which of the following pairs consists of an amphiprotic species and one of its possible roles?

A. HCl — acts as a base by donating H^+

B. HCO_3^- — acts as an acid by donating H^+ to form CO_3^{2-}

C. $NaOH$ — acts as an acid in some reactions

D. NH_4^+ — acts as a base by accepting H^+

Q51.

Q51. A buffer solution is prepared by mixing ethanoic acid and sodium ethanoate. A small amount of HCl is added to this buffer. Which equation correctly represents how the buffer responds?

A. $\mathrm{CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O}$

B. $\mathrm{CH_3COO^- + H^+ \rightarrow CH_3COOH}$

C. $\mathrm{CH_3COOH + H^+ \rightarrow CH_3COO^- + 2H^+}$

D. $\mathrm{Na^+ + Cl^- \rightarrow NaCl}$

Q52.

Worked Example: Finding Oxidation Numbers

What is the oxidation number of sulfur in $\mathrm{H_2SO_4}$?

Step 1: H = $+1$ each ($\times 2 = +2$); O = -2 each ($\times 4 = -8$)

Step 2: Sum must equal 0 (neutral compound): $+2 + x + (-8) = 0$

Step 3: $x = +6$

Q53.

Worked Example: Oxidation Number in a Polyatomic Ion

What is the oxidation number of Cr in $\mathrm{Cr_2O_7^{2-}}$?

Step 1: O = -2 each ($\times 7 = -14$)

Step 2: Sum must equal the ion charge (-2): $2x + (-14) = -2$

Step 3: $2x = 12$, so $x = +6$

Q54.

Example — Redox:

$\mathrm{2Mg + O_2 \rightarrow 2MgO}$

- Mg: $0 \rightarrow +2$ (oxidized — lost 2 electrons per atom)
- O: $0 \rightarrow -2$ (reduced — gained 2 electrons per atom)
- This IS a redox reaction. Mg is the reducing agent; $\mathrm{O_2}$ is the oxidizing agent.

Q55.

Example — NOT redox:

$\mathrm{HCl + NaOH \rightarrow NaCl + H_2O}$

- H: $+1 \rightarrow +1$; Cl: $-1 \rightarrow -1$; Na: $+1 \rightarrow +1$; O: $-2 \rightarrow -2$

- No oxidation numbers change. This is an acid-base (neutralization) reaction, NOT a redox reaction.

Q56.

Worked Example: Writing Half Equations

For the reaction: $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$

Oxidation half equation: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$

Reduction half equation: $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$

Check: Electrons lost (2) = electrons gained (2) ✓

Q57.

Worked Example: Balancing a Complex Half Equation

Write the balanced half equation for the reduction of MnO_4^- to Mn^{2+} in acidic solution.

Step 1: Mn is already balanced: $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$

Step 2: Balance O with water: $\text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

Step 3: Balance H with H^+: $\text{MnO}_4^- + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

Step 4: Balance charge: Left = $(-1) + (+8) = +7$. Right = $+2$. Add $5e^-$ to the left:

$\text{MnO}_4^- + 8\text{H}^+ + 5e^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

Check: Charge: $-1 + 8 - 5 = +2$ ✓ Atoms: Mn 1=1, O 4=4, H 8=8 ✓

Q58.

Worked Example: Calculating Cell EMF HL

Given: $E^\circ(\text{Cu}^{2+}/\text{Cu}) = +0.34$ V and $E^\circ(\text{Zn}^{2+}/\text{Zn}) = -0.76$ V

Which metal is the anode? Calculate the cell EMF.

Step 1: Zn has the more negative E°, so Zn is more easily oxidized. Zn is the anode. Cu is the cathode.

Step 2: $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = (+0.34) - (-0.76) = +1.10$ V

Q59.

Q59. In the reaction $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$, which substance is the reducing agent?

A. Fe_2O_3

B. CO

C. Fe

D. CO_2

Q60.

Q60. What is the oxidation number of nitrogen in NO_3^-?

A. -3

B. $+3$

C. $+5$

D. -5

Q61.

Q61. A strip of zinc is placed into a solution of $AgNO_3$. What is observed?

A. No reaction occurs

B. Silver dissolves into solution

C. The solution turns blue

D. A grey deposit forms on the zinc

Q62.

Q62. In a voltaic cell, oxidation occurs at the:

A. Cathode

B. Anode

C. Salt bridge

D. External wire

Q63.

Q63. What is the oxidation number of sulfur in H_2SO_4?

A. $+2$

B. $+4$

C. $+6$

D. -2

Q64.

Q64. In a voltaic (galvanic) cell using zinc and copper electrodes in their respective sulfate solutions, which statement is correct?

A. Copper is oxidized at the anode; zinc is reduced at the cathode

B. Zinc is oxidized at the anode; copper ions are reduced at the cathode

C. The salt bridge transfers electrons between the two half-cells

D. The cell produces a positive cell voltage only if copper is the more reactive metal

Q65.

Q65. Iron rusts according to the overall equation $4\mathrm{Fe} + 3\mathrm{O_2} + 6\mathrm{H_2O} \rightarrow 4\mathrm{Fe(OH)_3}$. What is the change in oxidation number of iron and the change in oxidation number of oxygen in this reaction?

A. Fe: 0 to +3 (oxidized); O: 0 to -2 (reduced)

B. Fe: +3 to 0 (reduced); O: -2 to 0 (oxidized)

C. Fe: 0 to +2 (oxidized); O: 0 to +2 (oxidized)

D. Fe: 0 to +3 (reduced); O: 0 to -2 (oxidized)

Q66.

Q66. Which compound belongs to the same homologous series as $\mathrm{CH_3CH_2CH_2OH}$?

A. $\mathrm{CH_3OCH_3}$

B. $\mathrm{CH_3CHO}$

C. $\mathrm{CH_3OH}$

D. $\mathrm{CH_3COOH}$

Q67.

Q67. Ethanal and propanone both have the general formula $\mathrm{C}_n\mathrm{H}_{2n}\mathrm{O}$. What is the correct term for the relationship between these two compounds?

A. Structural isomers

B. Homologues

C. Functional group isomers

D. Members of the same homologous series

Q68.

Q68. What is the correct IUPAC name for $CH_3CH(CH_3)CH_2CH_3$?

A. 3-methylbutane

B. 2-ethylpropane

C. 2-methylbutane

D. 1,1-dimethylpropane

Q69.

Q69. An organic compound has the structure $CH_3CH_2CH(OH)CH_3$. What is its IUPAC name?

A. 2-methylpropan-1-ol

B. Butan-2-ol

C. Butan-3-ol

D. 2-butanol

Q70.

Worked Example: Drawing isomers of C_5H_{12} systematically

Step 1: Start with the longest possible chain — 5 carbons in a row: $CH_3CH_2CH_2CH_2CH_3$ (pentane).

Step 2: Shorten the chain to 4 carbons, place the remaining carbon as a branch. Only one valid position (carbon 2): $CH_3CH(CH_3)CH_2CH_3$ (2-methylbutane). Putting the branch on carbon 1 just extends the chain back to pentane.

Step 3: Shorten the chain to 3 carbons, place 2 remaining carbons as branches. Both must go on the middle carbon: $C(CH_3)_4$ (2,2-dimethylpropane). That is all — 3 isomers total.

Q71.

Q71. How many structural isomers exist with the molecular formula C_4H_{10}?

A. 1

B. 2

C. 3

D. 4

Q72.

Q72. 2-Methylpropan-2-ol is classified as a tertiary alcohol. What does this mean?

A. The -OH group is on the third carbon in the chain

B. The molecule has three -OH groups

C. The molecule has three carbon atoms in total

D. The carbon bonded to -OH is attached to three other carbon atoms

Q73.

Q73. Which products are formed during the incomplete combustion of propane?

A. CO_2 and H_2O only

B. CO_2, CO, and H_2

C. CO, C, and H_2O

D. CO and H_2 only

Q74.

Q74. What is the balanced equation for the complete combustion of butane (C_4H_{10})?

A. $C_4H_{10} + 4O_2 \rightarrow 4CO_2 + 5H_2O$

B. $C_4H_{10} + 5O_2 \rightarrow 4CO_2 + 5H_2O$

C. $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

D. $C_4H_{10} + 6O_2 \rightarrow 4CO_2 + 4H_2O$

Q75.

Q75. Bromine water is added to a sample of an unknown hydrocarbon. The solution changes from orange to colourless. What can be concluded?

A. The compound is an alkane

B. The compound contains a -OH group

C. The compound contains a C=C double bond

D. The compound is saturated

Q76.

Q76. What is the product when propene ($CH_3CH{=}CH_2$) reacts with H_2O in the presence of H_3PO_4?

A. Propanal

B. Propan-2-ol

C. Propan-1-ol

D. Propanoic acid

Q77.

Q77. But-2-ene reacts with HBr. How many distinct organic products are formed?

A. 1

B. 2

C. 3

D. 4

Q78.

Worked Example: Proving a Primary Alcohol Has Been Oxidised to a Carboxylic Acid

Scenario: You heat ethanol with excess acidified $K_2Cr_2O_7$ *under reflux. How do you prove the product is ethanoic acid (not just that oxidation occurred)?*

Step	Test & Interpretation
Step 1: Prove oxidation occurred	Add acidified $K_2Cr_2O_7$ → solution turns orange to green. This proves electrons were transferred (oxidation happened).
Step 2: Prove it's not still an alcohol	Test with $K_2Cr_2O_7$: if orange goes green, something oxidisable is still present. If no more colour change, alcohol has been used up.
Step 3: Prove it's not an aldehyde	Add Tollens' reagent: if NO silver mirror forms, the product is NOT an aldehyde. This rules out ethanal.
Step 4: Prove it is a carboxylic acid	Test with litmus/pH paper: acidic pH (turns red) confirms a carboxylic acid is present. Add sodium carbonate solution: if CO_2 bubbles form, a carboxylic acid is confirmed (carboxylic acids react with Na_2CO_3 to release CO_2).
Conclusion	Orange → green (oxidation occurred) + no silver mirror (not aldehyde) + acidic pH + CO_2 with Na_2CO_3 → confirms ethanoic acid (carboxylic acid) has formed.

Q79.

Q79. An unknown alcohol is heated with acidified $K_2Cr_2O_7$. The solution stays orange. What type of alcohol is the unknown compound?

A. Primary

B. Secondary

C. Tertiary

D. It is not an alcohol

Q80.

Q80. What are the conditions needed to oxidise ethanol to ethanoic acid (rather than stopping at ethanal)?

A. Limited $K_2Cr_2O_7/H_2SO_4$, distillation

B. $NaBH_4$ in water, reflux

C. Excess $K_2Cr_2O_7/H_2SO_4$, reflux

D. Excess $K_2Cr_2O_7/H_2SO_4$, distillation

Q81.

Q81. Tollens' reagent is added to the product of oxidising propan-1-ol with limited $K_2Cr_2O_7/H_2SO_4$. What is observed?

A. No change — solution stays colourless

B. Blue solution turns brick-red

C. A silver mirror forms on the inside of the test tube

D. The solution turns green

Q82.

Example 1 — Reduction of ethanal to ethanol (aldehyde → primary alcohol):

$CH_3CHO + 2[H] \rightarrow CH_3CH_2OH$

ethanal → ethanol

Reagent: $NaBH_4$ in water. What happened: the C=O double bond became C-OH (added 2H). Reaction type: reduction.

Notice: this is the exact REVERSE of oxidising ethanol to ethanal with $K_2Cr_2O_7$.

Example 2 — Reduction of propanone to propan-2-ol (ketone → secondary alcohol):

$CH_3COCH_3 + 2[H] \rightarrow CH_3CH(OH)CH_3$

propanone → propan-2-ol

Reagent: $NaBH_4$ in water. What happened: C=O in middle of chain became C-OH. Reaction type: reduction.

The OH group ends up on the middle carbon — so the product is a SECONDARY alcohol.

Example 3 — Reduction of butanal to butan-1-ol:

$\mathrm{CH_3CH_2CH_2CHO + 2[H] \rightarrow CH_3CH_2CH_2CH_2OH}$

butanal → butan-1-ol

Reagent: $\mathrm{NaBH_4}$ in water. Type: reduction (aldehyde → primary alcohol).

The CHO at the end becomes $\mathrm{CH_2OH}$ *— the carbon count stays the same.*

Example 4 — Reduction of butanone to butan-2-ol:

$\mathrm{CH_3COCH_2CH_3 + 2[H] \rightarrow CH_3CH(OH)CH_2CH_3}$

butanone → butan-2-ol

Reagent: $\mathrm{NaBH_4}$. Type: reduction (ketone → secondary alcohol).

The C=O on carbon 2 becomes CHOH on carbon 2 — secondary alcohol.

Q83.

Q83. What is the product when butanone ($\mathrm{CH_3COCH_2CH_3}$) is reduced with $\mathrm{NaBH_4}$?

A. Butan-1-ol

B. Butan-2-ol

C. Butanal

D. Butanoic acid

Q84.

Q84. Which statement correctly describes the reduction of an aldehyde?

A. $\mathrm{K_2Cr_2O_7/H_2SO_4}$ is added and the solution turns green

B. [O] is added to convert C=O to C-OH

C. [H] is added to convert C=O to C-OH, using $\mathrm{NaBH_4}$ as the reducing agent

D. The aldehyde is heated under reflux to form a carboxylic acid

Q85.

Example 1 — Ethanoic acid + Ethanol:

$\mathrm{CH_3COOH + CH_3CH_2OH \xrightarrow{conc.\ H_2SO_4} CH_3COOCH_2CH_3 + H_2O}$

- Acid: ethanoic acid → provides the **ethanoate** part
- Alcohol: ethanol → provides the **ethyl** part
- Product: **ethyl ethanoate** (sweet, fruity smell — similar to nail polish remover)
- Conditions: concentrated H_2SO_4 catalyst, heat under **reflux**
- Reaction type: **condensation** (water is released)

Q86.

Example 2 — Methanoic acid + Propan-1-ol:

$$HCOOH + CH_3CH_2CH_2OH \xrightarrow{\text{conc. } H_2SO_4} HCOOCH_2CH_2CH_3 + H_2O$$

- Acid: methanoic acid → provides the **methanoate** part
- Alcohol: propan-1-ol → provides the **propyl** part
- Product: **propyl methanoate**

Q87.

Q87. Ethanoic acid reacts with methanol in the presence of concentrated H_2SO_4. What is the name of the ester produced?

A. Ethyl methanoate

B. Methyl ethanoate

C. Methyl methanoate

D. Ethyl ethanoate

Q88.

Q88. What type of reaction is esterification?

A. Addition

B. Substitution

C. Condensation

D. Elimination

Q89.

Example 1 — 1-Bromopropane + NaOH (SN2 — primary):

$$CH_3CH_2CH_2Br + NaOH(aq) \rightarrow CH_3CH_2CH_2OH + NaBr$$

1-bromopropane → propan-1-ol + sodium bromide

Mechanism: SN2 — one-step, OH^- attacks from the back simultaneously as Br^- leaves. Evidence: primary halogenoalkane = SN2. Rate depends on BOTH [RX] and [OH^-].

Example 2 — 2-Bromo-2-methylpropane + NaOH (SN1 — tertiary):

$(CH_3)_3CBr + NaOH(aq) \rightarrow (CH_3)_3COH + NaBr$

2-bromo-2-methylpropane → 2-methylpropan-2-ol

Mechanism: SN1 — two steps:

Step 1 (slow, rate-determining): $(CH_3)_3CBr \rightarrow (CH_3)_3C^+ + Br^-$

Step 2 (fast): $(CH_3)_3C^+ + OH^- \rightarrow (CH_3)_3COH$

Example 3 — Chloroethane + NaOH (SN2 — primary):

$CH_3CH_2Cl + NaOH(aq) \rightarrow CH_3CH_2OH + NaCl$

chloroethane → ethanol + sodium chloride

Note: C-Cl reacts more slowly than C-Br or C-I. Reactivity order: C-I > C-Br > C-Cl > C-F.

Example 4 — Bromoethane + KCN (different nucleophile — extends carbon chain):

$CH_3CH_2Br + KCN \rightarrow CH_3CH_2CN + KBr$

bromoethane → propanenitrile

Nucleophile: CN^- (cyanide ion). Replaces Br^-. Important: this EXTENDS the carbon chain by 1 carbon (C2 → C3 chain).

Q90.

Worked Example: Determining SN1 vs SN2 Mechanism

Problem: 2-chloro-2-methylbutane reacts with aqueous sodium hydroxide. (a) Classify the halogenoalkane. (b) Predict the mechanism. (c) Draw the mechanism with curly arrows. (d) Name the organic product.

Step 1: Classify the halogenoalkane

Draw the structure. The carbon bonded to Cl is bonded to THREE other carbons: CH_3, CH_3, and CH_2CH_3. This is a **TERTIARY** halogenoalkane.

Step 2: Predict the mechanism

Tertiary halogenoalkanes favor the **SN1 mechanism** (two-step, via carbocation intermediate). Tertiary carbocations are stabilized by three electron-donating alkyl groups, making them sufficiently stable to exist as intermediates.

Step 3: Draw the mechanism with curly arrows

Step 1 (slow, rate-determining):

$(CH_3)_2(C_2H_5)CCl \rightarrow (CH_3)_2(C_2H_5)C^+ + Cl^-$

Curly arrow: from the C–Cl bond to Cl (heterolytic fission — both electrons go to Cl, forming Cl^- and leaving behind a carbocation).

Step 2 (fast):

$(CH_3)_2(C_2H_5)C^+ + OH^- \rightarrow (CH_3)_2(C_2H_5)COH$

Curly arrow: from an OH^- lone pair to the carbocation (nucleophile donates electron pair to form C–O bond).

Step 4: Name the organic product

Product: **2-methylbutan-2-ol** (a tertiary alcohol).

Key takeaway: Always count the carbons attached to the C–X carbon. 1 carbon attached = primary = SN2. 3 carbons attached = tertiary = SN1. Tertiary carbocations are stabilized by hyperconjugation from three alkyl groups donating electron density.

IB mark scheme note: [3 marks] — 1 mark for correct classification (tertiary), 1 mark for mechanism choice with reasoning (SN1 because tertiary carbocation is stable), 1 mark for correct curly arrows or product structure.

Q91.

Worked Example: SN1 vs SN2 — Side-by-Side Mechanism Comparison

Problem: Compare the reactions of (A) 1-bromobutane with aqueous NaOH and (B) 2-bromo-2-methylpropane (tert-butyl bromide) with aqueous NaOH. For each, classify the substrate, predict the mechanism, draw the full mechanism with curly arrows, and explain why that mechanism is favored.

Reaction A: 1-bromobutane + NaOH(aq) → SN2 mechanism

Step 1: Classify the substrate

Structure: $CH_3CH_2CH_2CH_2Br$

The carbon bonded to Br is bonded to ONE other carbon ($CH_2CH_2CH_3$) and two hydrogens. This is a **PRIMARY** halogenoalkane.

Step 2: Predict the mechanism

Primary halogenoalkanes favor the **SN2 mechanism** (one-step, concerted). Reasons:

- Primary carbocations are extremely unstable (only one alkyl group for stabilization), so SN1 is not viable.
- The carbon bearing Br has minimal steric hindrance — the nucleophile can easily approach from the backside.

Step 3: Draw the mechanism with curly arrows

One-step mechanism (concerted):

$$OH^- + CH_3CH_2CH_2CH_2Br \rightarrow CH_3CH_2CH_2CH_2OH + Br^-$$

Curly arrows:

1. From a lone pair on OH^- to the carbon bearing Br (nucleophile attacks from backside).
2. From the C–Br bond to Br (C–Br bond breaks, both electrons go to Br, forming Br^-).

Transition state: The nucleophile approaches from the opposite side of the leaving group. At the transition state, the C–OH bond is partially formed and the C–Br bond is partially broken. The carbon center undergoes inversion of configuration (Walden inversion).

Step 4: Name the product

Product: **butan-1-ol** (a primary alcohol).

Why SN2? Primary substrate + minimal steric hindrance + strong nucleophile (OH^-) = SN2 pathway.

Reaction B: 2-bromo-2-methylpropane + NaOH(aq) → SN1 mechanism

Step 1: Classify the substrate

Structure: $(CH_3)_3CBr$ (tert-butyl bromide)

The carbon bonded to Br is bonded to THREE other carbons (three CH_3 groups). This is a **TERTIARY** halogenoalkane.

Step 2: Predict the mechanism

Tertiary halogenoalkanes favor the **SN1 mechanism** (two-step, via carbocation intermediate). Reasons:

- Tertiary carbocations are stabilized by three electron-donating alkyl groups (hyperconjugation and inductive effect).
- The carbon bearing Br is heavily sterically hindered — the nucleophile cannot easily approach for an SN2 backside attack.

Step 3: Draw the mechanism with curly arrows

Step 1 (slow, rate-determining):

Heterolytic fission of the C–Br bond:

$$(CH_3)_3CBr \rightarrow (CH_3)_3C^+ + Br^-$$

Curly arrow: from the C–Br bond to Br (both electrons go to Br, forming Br^- and leaving behind a tertiary carbocation $(CH_3)_3C^+$).

Carbocation intermediate: The tertiary carbocation $(CH_3)_3C^+$ is planar (sp^2 hybridized). It is stabilized by hyperconjugation from the three adjacent CH_3 groups.

Step 2 (fast):

Nucleophilic attack on the carbocation:

$$(CH_3)_3C^+ + OH^- \rightarrow (CH_3)_3COH$$

Curly arrow: from a lone pair on OH^- to the positively charged carbon (nucleophile donates electron pair to form the C–O bond).

Step 4: Name the product

Product: **2-methylpropan-2-ol** (a tertiary alcohol, also called tert-butanol).

Why SN1? Tertiary substrate + stable carbocation intermediate + significant steric hindrance = SN1 pathway.

Side-by-side summary:

Feature	Reaction A (SN2)	Reaction B (SN1)
Substrate classification	Primary (1-bromobutane)	Tertiary (2-bromo-2-methylpropane)
Number of steps	One (concerted)	Two (stepwise)
Intermediate	Transition state (no stable intermediate)	Carbocation intermediate $(CH_3)_3C^+$
Rate-determining step	Single step (bimolecular)	Step 1: C–Br bond breaking (unimolecular)
Stereochemistry	Inversion of configuration (Walden inversion)	Racemization (nucleophile can attack from either face of planar carbocation)
Why this mechanism?	Primary carbocation too unstable; minimal steric hindrance allows backside attack	Tertiary carbocation stable; steric hindrance prevents SN2

Key IB exam point: The classification (primary vs secondary vs tertiary) determines the mechanism. Always count the carbons attached to the C–X carbon FIRST. This tells you which mechanism is favored before you even draw anything.

IB mark scheme note: [4-5 marks total for a two-part comparison] — 1 mark for correct classification of each substrate, 1 mark for correct mechanism choice with reasoning (SN2 for primary, SN1 for tertiary), 2-3 marks for correct curly arrows and intermediate/transition state description.

Q92.

Q92. 2-Bromo-2-methylpropane reacts with aqueous NaOH. Which mechanism does this reaction follow?

A. SN2 — one-step, back-side attack

B. SN1 — two steps, via a carbocation intermediate

C. Free radical substitution

D. Electrophilic addition

Q93.

Q93. In an SN2 reaction, the rate depends on the concentration of:

A. Only the halogenoalkane

B. Only the nucleophile

C. Both the halogenoalkane and the nucleophile

D. Neither — it depends only on temperature

Q94.

Q94. In the free radical substitution of methane with chlorine, what is the initiation step?

A. $\mathrm{CH_4 \rightarrow CH_3\cdot + H\cdot}$

B. $\mathrm{Cl\cdot + CH_4 \rightarrow CH_3\cdot + HCl}$

C. $\mathrm{Cl_2 \rightarrow 2Cl\cdot}$

D. $\mathrm{CH_3\cdot + Cl\cdot \rightarrow CH_3Cl}$

Q95.

Q95. Which of the following is a termination step in the free radical substitution of methane with chlorine?

A. $\mathrm{Cl_2 \xrightarrow{UV} 2Cl\cdot}$

B. $\mathrm{Cl\cdot + CH_4 \rightarrow HCl + CH_3\cdot}$

C. $\mathrm{CH_3\cdot + Cl_2 \rightarrow CH_3Cl + Cl\cdot}$

D. $\mathrm{CH_3\cdot + CH_3\cdot \rightarrow C_2H_6}$

Q96.

Q96. What is the repeating unit of the polymer formed from chloroethene ($\mathrm{CH_2{=}CHCl}$)?

A. $[-CH_2=CHCl-]_n$

B. $[-CH_2\text{-}CH_2\text{-}CHCl-]_n$

C. $[-CH_2\text{-}CHCl-]_n$

D. $[-CHCl\text{-}CHCl-]_n$

Q97.

Q97. Which statement correctly distinguishes addition polymerization from condensation polymerization?

A. Addition polymers produce water as a by-product

B. Condensation polymers require monomers with C=C double bonds

C. Condensation polymerization releases a small molecule (e.g. H_2O) whereas addition polymerization does not

D. Addition polymerization uses bifunctional monomers

Q98.

Q98. Ethanol (CH_3CH_2OH) has a significantly higher boiling point than ethane (C_2H_6), despite having a similar molecular mass. What is the best explanation?

A. Ethanol has stronger covalent bonds

B. Ethanol has a higher molecular mass

C. Ethanol can form hydrogen bonds between molecules, whereas ethane can only form London dispersion forces

D. Ethanol is an ionic compound

Q99.

Question 1 [Functional Groups and Naming] [~6 marks]

Consider the following organic compound: $CH_3CH_2CH(OH)CH_2COOH$

(a) Identify the two functional groups present in this molecule.

(b) Name this compound using IUPAC nomenclature.

(c) State and explain one chemical test that could distinguish this compound from $CH_3CH_2CH_2CH_2CH_3$.

▶ Show Solution

Q100.

Question 2 [Nucleophilic Substitution] [~7 marks]

Halogenoalkanes undergo nucleophilic substitution reactions. Consider 1-bromobutane ($\mathrm{CH_3CH_2CH_2CH_2Br}$) and 2-bromo-2-methylpropane ($\mathrm{(CH_3)_3CBr}$).

(a) Define the term "nucleophile" and give two examples of nucleophiles.

(b) Explain why 1-bromobutane undergoes an $\mathrm{S_N2}$ mechanism while 2-bromo-2-methylpropane undergoes an $\mathrm{S_N1}$ mechanism.

(c) Outline the steps of the $\mathrm{S_N1}$ mechanism for the hydrolysis of 2-bromo-2-methylpropane with aqueous NaOH.

▶ Show Solution

Q101.

Question 3 [Reactions of Organic Compounds] [~6 marks]

Ethanol ($\mathrm{CH_3CH_2OH}$) can undergo several different types of reactions depending on conditions.

(a) Write a balanced equation for the complete combustion of ethanol.

(b) Ethanol can be oxidised to ethanal and then to ethanoic acid. State the reagent and conditions for each oxidation, and identify the type of organic compound formed at each stage.

(c) Ethanol reacts with ethanoic acid in the presence of a catalyst. Name the type of reaction, write the structural formula of the organic product, and name it.

▶ Show Solution

Q102.

Worked Example 1 — Mole and particle calculations

Part (a): Calculate the number of molecules in 4.40 g of carbon dioxide ($\mathrm{CO_2}$).

Step 1 — Molar mass: $M(\mathrm{CO_2}) = 12.01 + 2(16.00) = 44.01\ \mathrm{g\ mol^{-1}}$

Step 2 — Find moles: $n = \frac{4.40}{44.01} = 0.0999 \approx 0.100\ \mathrm{mol}$

Step 3 — Find number of molecules: $N = 0.100 \times 6.022 \times 10^{23} = 6.02 \times 10^{22}$ molecules

Part (b): What mass of sodium chloride (NaCl) contains 1.81×10^{24} formula units?

Step 1 — Find moles: $n = \dfrac{1.81 \times 10^{24}}{6.022 \times 10^{23}} = 3.005 \approx 3.00\ \mathrm{mol}$

Step 2 — Find mass: $M(\mathrm{NaCl}) = 22.99 + 35.45 = 58.44\ \mathrm{g\ mol^{-1}}$ $m = 3.00 \times 58.44 = 175.3 \approx 175\ \mathrm{g}$

Q103.

Worked Example 2 — Stoichiometric mass calculation

Magnesium burns in oxygen according to:

$$2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$$

What mass of MgO is produced when 3.60 g of Mg is completely burned?

Step 1: Find moles of Mg. $M(\text{Mg}) = 24.31\,\text{g mol}^{-1}$ $n(\text{Mg}) = \frac{3.60}{24.31} = 0.1481\,\text{mol}$

Step 2: Apply the mole ratio. From the equation, the ratio $\text{Mg} : \text{MgO} = 2 : 2 = 1 : 1$.
$n(\text{MgO}) = 0.1481\,\text{mol}$

Step 3: Find mass of MgO. $M(\text{MgO}) = 24.31 + 16.00 = 40.31\,\text{g mol}^{-1}$
$m(\text{MgO}) = 0.1481 \times 40.31 = 5.97\,\text{g}$

Q104.

Worked Example 3 — Finding the limiting reagent

Iron reacts with chlorine gas:

$$2\text{Fe}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{FeCl}_3(s)$$

15.0 g of Fe is mixed with 15.0 g of Cl_2. Identify the limiting reagent and calculate the mass of FeCl_3 produced.

Step 1: Convert to moles. $n(\text{Fe}) = \frac{15.0}{55.85} = 0.2686\,\text{mol}$ $n(\text{Cl}_2) = \frac{15.0}{70.90} = 0.2116\,\text{mol}$

Step 2: Divide by stoichiometric coefficients to compare. $\frac{n(\text{Fe})}{2} = \frac{0.2686}{2} = 0.1343$ $\frac{n(\text{Cl}_2)}{3} = \frac{0.2116}{3} = 0.0705$

The smaller value is for Cl_2, so Cl_2 **is the limiting reagent**.

Step 3: Calculate moles of FeCl_3 based on limiting reagent. Ratio $\text{Cl}_2 : \text{FeCl}_3 = 3 : 2$ $n(\text{FeCl}_3) = 0.2116 \times \frac{2}{3} = 0.1411\,\text{mol}$

Step 4: Find mass of FeCl_3. $M(\text{FeCl}_3) = 55.85 + 3(35.45) = 162.20\,\text{g mol}^{-1}$
$m(\text{FeCl}_3) = 0.1411 \times 162.20 = 22.9\,\text{g}$

Q105.

Worked Example 4 — Percentage yield

In the reaction: $\text{CH}_3\text{COOH}(l) + \text{C}_2\text{H}_5\text{OH}(l) \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5(l) + \text{H}_2\text{O}(l)$

12.0 g of ethanoic acid (CH_3COOH) is reacted with excess ethanol. 10.4 g of ethyl ethanoate ($\text{CH}_3\text{COOC}_2\text{H}_5$) is isolated.

Calculate the percentage yield.

Step 1: Find moles of $\mathrm{CH_3COOH}$ (the limiting reagent, since ethanol is in excess).
$M(\mathrm{CH_3COOH}) = 2(12.01) + 4(1.008) + 2(16.00) = 60.05\,\mathrm{g\ mol^{-1}}$
$n(\mathrm{CH_3COOH}) = \frac{12.0}{60.05} = 0.1998\,\mathrm{mol}$

Step 2: Find theoretical yield of ester. Mole ratio $1:1$, so $n(\mathrm{ester}) = 0.1998\,\mathrm{mol}$.
$M(\mathrm{CH_3COOC_2H_5}) = 4(12.01) + 8(1.008) + 2(16.00) = 88.11\,\mathrm{g\ mol^{-1}}$
$\mathrm{theoretical\ yield} = 0.1998 \times 88.11 = 17.60\,\mathrm{g}$

Step 3: Calculate percentage yield. $\%\ \mathrm{yield} = \frac{10.4}{17.60} \times 100 = 59.1\%$

Q106.

Worked Example 5 — Atom economy

Chlorine gas can be produced by two processes:

Process 1 (addition/synthesis type): $\mathrm{2NaCl}(s) + \mathrm{2H_2O}(l) \rightarrow \mathrm{Cl_2}(g) + \mathrm{H_2}(g) + \mathrm{2NaOH}(aq)$

Here the desired product is $\mathrm{Cl_2}$. $\mathrm{H_2}$ and NaOH are co-products (may be sold or used).

$M(\mathrm{Cl_2}) = 70.90\,\mathrm{g\ mol^{-1}}$ $\mathrm{Total\ molar\ mass\ of\ all\ products} = 70.90 + 2.016 + 2(40.00) = 152.92\,\mathrm{g\ mol^{-1}}$ $\%\ \mathrm{atom\ economy} = \frac{70.90}{152.92} \times 100 = 46.4\%$

Note: If both $\mathrm{Cl_2}$ and NaOH are considered desired products (both are commercially valuable), the atom economy increases substantially — context matters.

Process 2 (substitution type, simplified): If only $\mathrm{Cl_2}$ is desired and all other products are waste, atom economy = 46.4% as calculated.