



Premium IB tuition — Physics, Chemistry, Maths

IB CHEMISTRY · STRUCTURE 1.4.5 – 1.4.6

Solutions and Gases: Concentration & Avogadro's Law

Notes · Formulae · Solved Examples · Practice & Mark Schemes

STANDARD LEVEL

HIGHER LEVEL

HL STRETCH INCLUDED

First assessment 2025 syllabus | Paper 1A (MCQ), Paper 1B (data-based) and Paper 2 practice

16 fully solved examples · 20 multiple-choice questions

5 exam-style questions with full mark schemes

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AN IMPORTANT NOTE ON SL AND HL FOR THIS SUB-TOPIC

In the current IB Chemistry syllabus (first assessment 2025), **the whole of Structure 1.4 is common core content** — studied and examined at both Standard Level and Higher Level. There is **no AHL (HL-only) material in 1.4.5 or 1.4.6**. Anyone who tells you otherwise is working from the old (2016) syllabus.

So that this booklet is still genuinely useful to HL students, every section is labelled as follows:

LABEL	WHAT IT MEANS
SL + HL	Examinable syllabus content for both levels. This is the great majority of the booklet — all notes, all formulae, both sub-topics.
HL STRETCH	Not extra syllabus content. Harder question styles, multi-step problems and forward links to AHL topics where HL candidates are expected to apply concentration and gas calculations with greater sophistication. SL students may safely skip these; strong SL candidates will benefit from attempting them.
S1.5	Content that formally belongs to the next sub-topic (Structure 1.5, Ideal gases), included here because 1.4.6 questions almost always need it. Also common core, SL and HL.

HOW TO GET THE MOST OUT OF THIS BOOKLET

- ▶ **Read Part A actively.** Cover the worked lines with a card and try each one before reading on.
- ▶ **Memorise Part B.** Only $n = cV$ and the molar volume are given in the IB data booklet — the dilution and gas-ratio relationships are not.
- ▶ **Do Part E under timed conditions.** Allow about 1.5 minutes per MCQ and roughly 1.5 minutes per mark for Papers 1B and 2. Under the 2025 syllabus a calculator *and* the data booklet are permitted in every paper, including Paper 1A — so the challenge is speed and accuracy, not mental arithmetic.
- ▶ **Mark yourself honestly against Part F,** then re-do every question you lost a mark on *from scratch* two days later.

Data used throughout: A_r values are quoted to 2 decimal places as given in the IB Chemistry data booklet periodic table; the Avogadro constant is taken as $N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$; the molar volume of an ideal gas at STP is taken as $22.7 \text{ dm}^3 \text{ mol}^{-1}$ (STP = 273 K and 100 kPa). Final answers are given to 3 significant figures unless stated otherwise.

PART A — NOTES & KEY POINTS

STRUCTURE 1.4.5

Molar concentration

SL + HL

Syllabus statement: the molar concentration is determined by the amount of solute and the volume of solution.

Why concentration matters

Sub-topics 1.4.1–1.4.4 let you convert between mass, amount and particle count for a *pure* substance. But most chemistry — every titration, every rate experiment, every buffer — happens in **solution**, where the substance you care about is dispersed through a solvent. You cannot weigh out the solute once it has dissolved. Concentration is the bridge that lets you measure a *volume* with a pipette and know exactly what amount of substance you have transferred.

DEFINITIONS YOU MUST KNOW

Solute — the substance that dissolves. **Solvent** — the liquid it dissolves in. **Solution** — the homogeneous mixture of the two.

Molar concentration (c) — the amount of solute, in mol, per unit volume of *solution*. Units: **mol dm⁻³**.

Standard solution — a solution whose concentration is known accurately, prepared by dissolving a weighed mass of solute and making the solution up to a precise total volume in a volumetric flask.

THE CENTRAL RELATIONSHIP — THIS ONE IS IN THE DATA BOOKLET

$$c = \frac{n}{V} \quad n = cV \quad V = \frac{n}{c}$$

c = concentration in mol dm⁻³ · n = amount of solute in mol · V = volume of **solution** in dm³

THE SINGLE MOST EXPENSIVE MISTAKE IN THIS WHOLE SUB-TOPIC

Volumes in the laboratory are measured in cm³. The relationship needs dm³. Burettes, pipettes and volumetric flasks are all marked in cm³, so almost every exam question hands you a number that must be divided by 1000 before you touch it. Convert on line one, every time.

$$1 \text{ dm}^3 = 1000 \text{ cm}^3 = 1 \text{ litre} \implies V/\text{dm}^3 = \frac{V/\text{cm}^3}{1000}$$

25.00 cm³ = 0.02500 dm³ · 250.0 cm³ = 0.2500 dm³ · 1.00 cm³ = 1.00 × 10⁻³ dm³. A factor-of-1000 slip is almost always this.

Square brackets — the notation the guide requires

Square brackets around a formula mean "the molar concentration of". You are expected to use and read this notation fluently:

$$[\text{NaCl}] = 0.100 \text{ mol dm}^{-3} \text{ means "the concentration of NaCl is } 0.100 \text{ mol dm}^{-3}\text{"}$$

Brackets are used only for *molar* concentration. Never write "[NaCl] = 5.85 g dm⁻³" — that is a mass concentration and takes no brackets.

Mass concentration: g dm⁻³

The syllabus requires you to convert between the two units of concentration. The bridge is, as always, the molar mass:

$$\text{concentration in g dm}^{-3} = \text{concentration in mol dm}^{-3} \times M \iff \text{mol dm}^{-3} = \frac{\text{g dm}^{-3}}{M}$$

Sense-check: multiplying by M (a number typically between 2 and 400) should make the number *bigger*. If your g dm⁻³ figure is smaller than your mol dm⁻³ figure, you have divided instead of multiplied.

SOLUTION	MOL DM ⁻³	M / G MOL ⁻¹	G DM ⁻³
Sodium chloride	0.100	58.44	5.84
Glucose	0.0500	180.18	9.01
Copper(II) sulfate	0.0200	159.62	3.19
Sodium hydroxide	0.250	40.00	10.0

Concentration of ions in solution

When an ionic compound dissolves it dissociates, and the concentration of each ion is the concentration of the compound multiplied by that ion's subscript. This is exactly the entity-counting skill from 1.4.1, applied in solution — and it is a reliable source of Paper 1A questions.

SOLUTION	DISSOCIATION	ION CONCENTRATIONS
0.10 mol dm ⁻³ NaCl	$\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$	$[\text{Na}^+] = 0.10 \cdot [\text{Cl}^-] = 0.10$
0.10 mol dm ⁻³ MgCl ₂	$\text{MgCl}_2 \rightarrow \text{Mg}^{2+} + 2\text{Cl}^-$	$[\text{Mg}^{2+}] = 0.10 \cdot [\text{Cl}^-] = 0.20$
0.10 mol dm ⁻³ Na ₂ SO ₄	$\text{Na}_2\text{SO}_4 \rightarrow 2\text{Na}^+ + \text{SO}_4^{2-}$	$[\text{Na}^+] = 0.20 \cdot [\text{SO}_4^{2-}] = 0.10$
0.10 mol dm ⁻³ Al ₂ (SO ₄) ₃	$\text{Al}_2(\text{SO}_4)_3 \rightarrow 2\text{Al}^{3+} + 3\text{SO}_4^{2-}$	$[\text{Al}^{3+}] = 0.20 \cdot [\text{SO}_4^{2-}] = 0.30$

Preparing a standard solution — the practical routine

- 1 **Calculate the mass needed:** $n = cV$, then $m = nM$.
- 2 **Weigh accurately** on an analytical balance, using the "weighing by difference" method.
- 3 **Dissolve fully in a beaker** in a volume of solvent well below the final volume.
- 4 **Transfer quantitatively** to a volumetric flask, rinsing the beaker and stirring rod into the flask so that no solute is left behind.
- 5 **Make up to the graduation mark** with solvent, reading the bottom of the meniscus at eye level, then stopper and invert repeatedly to mix.

GLASSWARE — A TOOL 1 SKILL THE GUIDE ASKS ABOUT EXPLICITLY

APPARATUS	USE IT FOR	TYPICAL UNCERTAINTY
Volumetric flask	Making a solution up to one exact total volume	$\pm 0.20 \text{ cm}^3$ (250 cm^3)
Volumetric pipette	Delivering one fixed, accurately known volume	$\pm 0.06 \text{ cm}^3$ (25 cm^3)
Burette	Delivering a variable, accurately measured volume	$\pm 0.05 \text{ cm}^3$ per reading
Measuring cylinder	Approximate volumes only — never for quantitative work	$\pm 0.5 \text{ cm}^3$ (25 cm^3)
Beaker	Dissolving and mixing — never for measuring	$\pm 5 \%$ or worse

Dilution

Adding solvent to a solution increases the volume but leaves the *amount* of solute unchanged. That single fact gives the dilution relationship:

Since n before = n after, and $n = cV$:

$$c_1 V_1 = c_2 V_2$$

Subscript 1 = before dilution (the concentrated stock), subscript 2 = after. Because the equation contains a volume on *both* sides, the units cancel — so this is the one relationship where you may safely leave both volumes in cm^3 , provided you use the same unit throughout.

DILUTION TRAPS

- ▶ **V_2 is the final total volume, not the volume of water added.** "25.0 cm³ is diluted to 250.0 cm³" means $V_2 = 250.0$. "25.0 cm³ is diluted *with* 250.0 cm³ of water" means $V_2 = 275.0$. Read the preposition.
- ▶ **Concentration is an *intensive* property.** Mixing two portions of the *same* solution does not change its concentration — pouring two beakers of 0.500 mol dm⁻³ NaOH together still gives 0.500 mol dm⁻³.
- ▶ **Dilution factor.** Diluting from 25.0 cm³ to 250.0 cm³ is a factor of 10, so the concentration falls by a factor of 10. Quote the factor in your working — it makes ECF easy for the examiner to award.
- ▶ **Serial dilutions multiply.** Three successive tenfold dilutions give a factor of 10³, not 30.

BEYOND THE GUIDE PARTS PER MILLION (PPM)

The current guide lists only g dm⁻³ and mol dm⁻³ for Structure 1.4.5, so **ppm is not required** for this sub-topic. It is worth knowing anyway, because it appears constantly in environmental data, in Reactivity topics and in IA work: ppm is the mass of solute per 10⁶ units of mass of solution. For dilute *aqueous* solutions, where 1 dm³ of solution has a mass of very nearly 1000 g, the useful shortcut is **1 ppm $\approx$ 1 mg dm⁻³**.

HL STRETCH WHERE CONCENTRATION REAPPEARS AT HL

$c = n/V$ is the entry point to a very large fraction of the AHL course: rate expressions and orders of reaction (Reactivity 2.2, where rate is expressed in mol dm⁻³ s⁻¹); the equilibrium constant K_c (Reactivity 2.3); pH, pK_a , buffers and the Henderson–Hasselbalch relationship (Reactivity 3.1), where every quantity in sight is a concentration; and electrode potentials (Reactivity 3.2). A student who is still converting cm³ to dm³ unreliably in Year 1 will lose marks in every one of those topics in Year 2.

STRUCTURE 1.4.6

Avogadro's law

SL + HL

Syllabus statement: Avogadro's law states that equal volumes of all gases measured under the same conditions of temperature and pressure contain equal numbers of molecules.

The idea behind the law

In a gas the molecules are separated by distances far larger than the molecules themselves, so almost all of the volume of a gas sample is empty space. The consequence is remarkable and slightly counter-intuitive: **the volume occupied by a gas does not depend on what the gas is**. A helium atom and a large propane molecule differ enormously in mass and in size, yet at the same temperature and pressure each one commands, on average, the same share of the container. Volume depends only on *how many* particles there are.

THE LAW, AND WHAT FOLLOWS FROM IT

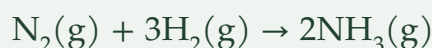
Avogadro's law — equal volumes of all gases, measured at the same temperature and pressure, contain equal numbers of molecules.

Equivalently, at constant temperature and pressure the volume of a gas is **directly proportional to the amount of gas**:

$$V \propto n \quad \text{or} \quad \frac{V_1}{n_1} = \frac{V_2}{n_2} \quad (\text{constant } T \text{ and } p)$$

The pay-off: for gases, volume ratio = mole ratio

This is what the law is *for* in an examination. If every substance in an equation is a gas and all volumes are measured at the same temperature and pressure, the coefficients of the balanced equation can be read directly as a ratio of **volumes**. You never need to calculate an amount in mol at all.

WORKED ILLUSTRATION — THE HABER PROCESS

	N ₂	H ₂	NH ₃
Mole ratio from the equation	1	3	2
Volume ratio (same <i>T, p</i>)	1	3	2
So 50 cm ³ of N ₂ reacts with...	50 cm ³	150 cm ³	100 cm ³

Notice that the total volume falls from 200 cm³ to 100 cm³. Gas volumes are *not* conserved in a reaction — only mass is.

THE FOUR CONDITIONS — ALL MUST HOLD BEFORE YOU USE VOLUME RATIOS

- ▶ **Everything must be a gas.** A state symbol of (s), (l) or (aq) removes that species from the volume argument entirely. In $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$, only the O₂ and CO₂ take part in the volume ratio.
- ▶ **Same temperature and pressure** for every measurement, including the measurement *after* the reaction.
- ▶ **Watch the water.** In a combustion question, water is a gas above 100 °C but a liquid at room temperature. "The mixture was cooled to the original temperature" is the examiner telling you the water has condensed and must be removed from your total.
- ▶ **Check for a limiting reactant** before assuming all of both gases reacts. Any excess reactant is still there afterwards and still occupies volume.

Gas volume problems — a reliable four-step method

- 1 Write the **balanced equation** with state symbols. Underline the gases.
- 2 Write each given volume under its **formula** and divide each by its coefficient. The *smallest* result identifies the limiting reactant.
- 3 Scale the **whole equation** to the limiting reactant to find the volume of every gas used and produced.
- 4 Add up what is **actually present at the end**: products that are gases at the stated temperature, plus any unreacted excess reactant.

NATURE OF SCIENCE — GAY-LUSSAC, AVOGADRO AND A FIFTY-YEAR ARGUMENT

In 1808 Gay-Lussac observed that gases combine in simple whole-number volume ratios. In 1811 Avogadro proposed the explanation, and in doing so had to make a claim that his contemporaries found unacceptable: that elements such as hydrogen and oxygen exist as *diatomic molecules*, H_2 and O_2 , rather than single atoms. Without that step the observed 2 : 1 : 2 ratio for water simply cannot be explained. Avogadro's hypothesis was largely ignored for almost fifty years until Cannizzaro revived it at the Karlsruhe Congress in 1860, after which it rapidly resolved the chaos in atomic masses. A good TOK example of a correct idea rejected because it required abandoning a prior assumption.

BRIDGE TO STRUCTURE 1.5.3

The molar volume of an ideal gas

S1.5 · SL + HL

Avogadro's law tells you that a given amount of *any* gas occupies the same volume under the same conditions. It does not tell you *what* that volume is. That number belongs formally to Structure 1.5.3 — but almost every question that converts between a mass and a gas volume needs it, so it is included here.

MOLAR VOLUME, V_M

Molar volume — the volume occupied by one mole of a gas at a specified temperature and pressure.

At STP (standard temperature and pressure: **273 K and 100 kPa**), the molar volume of an ideal gas is **$22.7 \text{ dm}^3 \text{ mol}^{-1}$** . This value *is* given in Section 2 of the IB data booklet, so you need not memorise it — but you must recognise when to reach for it.

$$n = \frac{V}{V_m} = \frac{V / \text{dm}^3}{22.7} \quad \Leftrightarrow \quad V = n \times 22.7 \text{ dm}^3 \quad (\text{at STP only})$$

A useful consequence: the density of a gas at STP is $M \div 22.7$ in g dm^{-3} . Rearranged, **$M = \text{density} \times 22.7$** — which is how an unknown gas is identified from a density measurement.

MOLAR VOLUME TRAPS

- ▶ **22.7, not 22.4.** The older figure of $22.4 \text{ dm}^3 \text{ mol}^{-1}$ is for 273 K and 101.3 kPa (1 atm). The IB defines STP with a pressure of 100 kPa, giving 22.7. Textbooks and websites written for other syllabuses will disagree with your data booklet — trust the data booklet.
- ▶ **Molar volume applies to gases only.** There is no molar volume for a solid or a liquid, and it never applies to a solution.
- ▶ **It is not room temperature.** STP is 273 K, which is 0°C . If a question specifies a different temperature or pressure, you need $pV = nRT$ from Structure 1.5.3 instead, with $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$, p in Pa, V in m^3 and T in K.
- ▶ **Watch the volume unit.** V_{m} is in $\text{dm}^3 \text{ mol}^{-1}$. An answer wanted in cm^3 needs a final $\times 1000$.

HL STRETCH WHERE GAS CALCULATIONS REAPPEAR AT HL

Gas volume ratios underpin: partial pressures and K_{p} (Reactivity 2.3, where the mole fraction of a gas equals its volume fraction — a direct consequence of Avogadro's law); Le Châtelier's principle applied to a change in pressure, which requires you to count moles of gas on each side of an equation; enthalpies of combustion; and atom economy calculations for gas-phase industrial processes. The forward link the guide itself poses is worth thinking about now: Avogadro's law holds exactly only for an *ideal* gas, so under what conditions — and for which gases — would you expect it to fail? (Answer: high pressure and low temperature, and most severely for gases with strong intermolecular forces such as NH_3 and H_2O .)

PART B — FORMULA & DATA SHEET

Everything on this page applies to **both SL and HL**. Only the items marked "given" appear in the IB data booklet — the rest must be memorised.

RELATIONSHIP	WHAT IT DOES	UNITS OF ANSWER	IN DATA BOOKLET?
$c = n / V$	Amount and volume of solution → concentration	mol dm^{-3}	Given (as $n = cV$)
$n = c \times V$	Concentration and volume → amount of solute	mol	Given
$V = n / c$	What volume delivers a required amount	dm^3	Given
$\text{g dm}^{-3} = \text{mol dm}^{-3} \times M$	Molar concentration → mass concentration	g dm^{-3}	No
$c_1 V_1 = c_2 V_2$	Dilution of a stock solution	matching units	No
$m = cVM$	Mass of solute to make a standard solution	g	No (combination)
$V \propto n$ (constant T, p)	Avogadro's law — volume ratio = mole ratio	—	No
$V_1/n_1 = V_2/n_2$	Change in gas volume on changing amount	matching units	No
$V_m = 22.7 \text{ dm}^3 \text{ mol}^{-1}$	Molar volume of an ideal gas at STP	$\text{dm}^3 \text{ mol}^{-1}$	Given
$n = V / 22.7$	Gas volume at STP → amount	mol	No
$M = \text{density} \times 22.7$	Gas density at STP → molar mass	g mol^{-1}	No
$pV = nRT$ S1.5.3	Any gas at any stated T and p	SI units throughout	Given

SYMBOLS, UNITS AND CONVERSIONS — LEARN THIS TABLE VERBATIM

SYMBOL	QUANTITY	UNIT	WATCH OUT FOR
c or $[X]$	Molar concentration	mol dm^{-3}	Square brackets are for molar concentration only.
V	Volume of solution	dm^3	Divide a cm^3 value by 1000 before use.
n	Amount of solute or gas	mol	Always state which substance.
V_m	Molar volume of a gas	$\text{dm}^3 \text{ mol}^{-1}$	22.7 at STP; gases only.
R	Ideal gas constant	$\text{J K}^{-1} \text{ mol}^{-1}$	8.31; forces SI units on p , V and T .

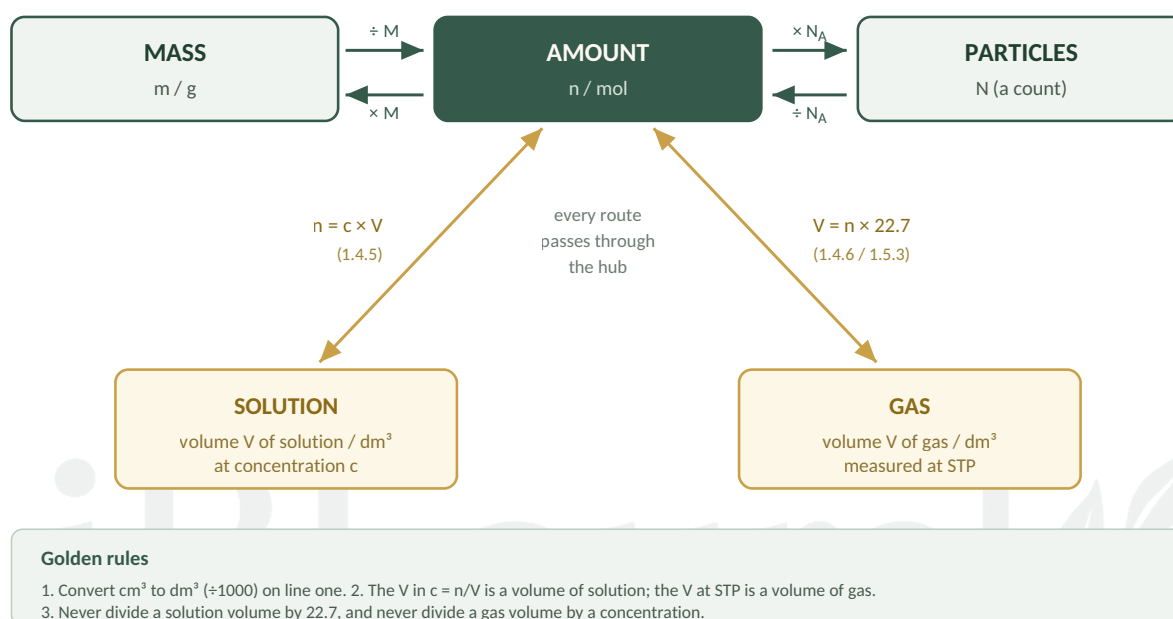
Conversions to have at your fingertips: $1 \text{ dm}^3 = 1000 \text{ cm}^3 = 1 \text{ L} = 1 \times 10^{-3} \text{ m}^3$ · $1 \text{ m}^3 = 1000 \text{ dm}^3$ · $1 \text{ kPa} = 1000 \text{ Pa}$ · $T/\text{K} = \theta/^{\circ}\text{C} + 273$

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PART C — EXAM TECHNIQUE

The solution and gas road map

Sub-topics 1.4.1–1.4.4 gave you the top row of this diagram. 1.4.5 and 1.4.6 add the two boxes underneath. Amount in mol remains the hub: every route in and out of a problem passes through it.



Choosing the right route — a decision checklist

THE QUESTION GIVES YOU...	...AND WANTS...	USE
Volume of solution + concentration	Amount of solute	$n = cV$, with V in dm ³
Mass of solute + volume of solution	Concentration	$n = m/M$, then $c = n/V$
A required concentration and volume	Mass to weigh out	$m = cVM$
Stock concentration and both volumes	Final concentration	$c_1V_1 = c_2V_2$
Two titre volumes and one concentration	The other concentration	$n = cV$ for the known, mole ratio, then $c = n/V$
Gas volumes only, same T and p	Another gas volume	Volume ratio = coefficient ratio. No moles needed.
A mass or amount, and STP is stated	A gas volume	$V = n \times 22.7$ dm ³
Gas volume at STP	Amount or mass	$n = V/22.7$, then $m = nM$
p, V and T that are <i>not</i> STP	Amount of gas	$pV = nRT$ in SI units S1.5.3

#	HABIT	WHY IT MATTERS
1	Convert cm^3 to dm^3 on line one	Removes the factor-of-1000 error that dominates this sub-topic before it can propagate.
2	Write the relationship before the numbers	Method marks are awarded for the correct relationship even when the arithmetic fails.
3	Label every line with the substance	" $n(\text{NaOH}) = 2.15 \times 10^{-3} \text{ mol}$ " is markable; a bare number may not be.
4	Write the balanced equation first for any titration	The mole ratio is where most marks are lost — 1 : 1 is an assumption, not a fact.
5	Check state symbols before using volume ratios	Only (g) species count. Solids, liquids and solutions do not occupy gas volume.
6	Ask "has the water condensed?"	The difference between the answer and a zero is often one line about temperature.
7	Test for a limiting reactant whenever two amounts are given	Divide each amount by its coefficient; the smallest wins. Excess reactant still occupies volume.
8	State the dilution factor explicitly	"Diluted by a factor of 10" is one clean line an examiner can follow and award.
9	Attach the unit to every answer	mol dm^{-3} , g dm^{-3} , dm^3 , cm^3 , mol — on Paper 2 the unit is often the mark.
10	Sanity-check the magnitude	Bench solutions are usually $0.01\text{--}2 \text{ mol dm}^{-3}$; laboratory gas volumes $10\text{--}1000 \text{ cm}^3$.
11	Never round mid-calculation	Titration answers to 3 s.f. are unforgiving of an early rounding.
12	Re-read the stem after answering	Confirms mol dm^{-3} or g dm^{-3} , dm^3 or cm^3 , before or after dilution.

VALUES WORTH KNOWING BY SIGHT

A calculator and the data booklet are permitted in every paper including Paper 1A, so the constraint is **speed**, not arithmetic. Recognise these instantly: $25.00 \text{ cm}^3 = 0.02500 \text{ dm}^3$; $250.0 \text{ cm}^3 = 0.2500 \text{ dm}^3$; a 25 cm^3 aliquot into a 250 cm^3 flask is a tenfold dilution; 1 mol of gas at STP = 22.7 dm^3 , so $0.100 \text{ mol} = 2.27 \text{ dm}^3$ and $0.0100 \text{ mol} = 227 \text{ cm}^3$; $M(\text{NaOH}) = 40.00$, so a 40.0 g dm^{-3} solution is exactly 1.00 mol dm^{-3} ; and for combustion of a hydrocarbon C_xH_y , the oxygen coefficient is always $x + y/4$.

PART D — FULLY SOLVED EXAMPLES

Cover the solution, attempt the question, then compare. All sixteen examples are **SL + HL** core content; Examples 9, 15 and 16 are additionally marked **HL Stretch** because of the extra reasoning step, not because the content differs by level.

Sub-topic 1.4.5 — Concentration and solutions

SL + HL

Example 1. 5.85 g of sodium chloride is dissolved in water and the solution made up to 250.0 cm³. Calculate the molar concentration of the solution.

- 1 **Convert the volume first:** $V = 250.0 \div 1000 = 0.2500 \text{ dm}^3$
- 2 $M(\text{NaCl}) = 22.99 + 35.45 = 58.44 \text{ g mol}^{-1}$
- 3 $n(\text{NaCl}) = m/M = 5.85 \div 58.44 = 0.10010 \text{ mol}$
- 4 $c = n/V = 0.10010 \div 0.2500 = 0.40041 \text{ mol dm}^{-3}$

[NaCl] = 0.400 mol dm⁻³

Had you left the volume as 250.0, you would have obtained 4.00×10^{-4} — a thousand times too small, and instantly implausible for a bench solution.

Example 2. Calculate the mass of anhydrous sodium carbonate, Na₂CO₃, needed to prepare 250.0 cm³ of a 0.100 mol dm⁻³ standard solution.

- 1 This runs the previous example backwards: $n = cV$, then $m = nM$.
- 2 $V = 0.2500 \text{ dm}^3$, so $n = 0.100 \times 0.2500 = 0.0250 \text{ mol}$
- 3 $M(\text{Na}_2\text{CO}_3) = 2(22.99) + 12.01 + 3(16.00) = 45.98 + 12.01 + 48.00 = 105.99 \text{ g mol}^{-1}$
- 4 $m = 0.0250 \times 105.99 = 2.6498 \text{ g}$

2.65 g

Note the practical realism: 2.65 g is a mass an analytical balance can measure to four significant figures. If your answer came out at 0.00265 g or 2650 g, a unit conversion has gone wrong.

Example 3. A glucose solution has $[\text{C}_6\text{H}_{12}\text{O}_6] = 0.0500 \text{ mol dm}^{-3}$. (a) Express this concentration in g dm⁻³. (b) A different solution contains 18.0 g dm⁻³ of glucose. Express this in mol dm⁻³.

- 1 $M(\text{C}_6\text{H}_{12}\text{O}_6) = 6(12.01) + 12(1.01) + 6(16.00) = 72.06 + 12.12 + 96.00 = 180.18 \text{ g mol}^{-1}$
- 2 (a) $\text{g dm}^{-3} = \text{mol dm}^{-3} \times M = 0.0500 \times 180.18 = 9.009 \text{ g dm}^{-3}$
- 3 (b) $\text{mol dm}^{-3} = \text{g dm}^{-3} \div M = 18.0 \div 180.18 = 0.09990 \text{ mol dm}^{-3}$

(a) 9.01 g dm⁻³ (b) 0.0999 mol dm⁻³

No volume appears anywhere in this calculation, because both quantities are already "per dm³". A common error is to invent a volume and multiply by it.

Example 4. Determine the concentration of each ion in a $0.150 \text{ mol dm}^{-3}$ solution of aluminium sulfate, $\text{Al}_2(\text{SO}_4)_3$.

- 1 Write the dissociation: $\text{Al}_2(\text{SO}_4)_3(\text{aq}) \rightarrow 2\text{Al}^{3+}(\text{aq}) + 3\text{SO}_4^{2-}(\text{aq})$
- 2 $[\text{Al}^{3+}] = 2 \times 0.150 = 0.300 \text{ mol dm}^{-3}$
- 3 $[\text{SO}_4^{2-}] = 3 \times 0.150 = 0.450 \text{ mol dm}^{-3}$

$$[\text{Al}^{3+}] = 0.300 \text{ mol dm}^{-3} \quad \cdot \quad [\text{SO}_4^{2-}] = 0.450 \text{ mol dm}^{-3}$$

The sulfate ion is a single polyatomic ion, so there are 3 sulfate ions per formula unit, not 12 oxygens' worth. Total ion concentration is $0.300 + 0.450 = 0.750 \text{ mol dm}^{-3}$.

Example 5. 25.0 cm^3 of 2.00 mol dm^{-3} hydrochloric acid is transferred to a volumetric flask and made up to 500.0 cm^3 with distilled water. Calculate the concentration of the diluted solution.

- 1 Use $c_1 V_1 = c_2 V_2$. Because volume appears on both sides, cm^3 may be used throughout.
- 2 $2.00 \times 25.0 = c_2 \times 500.0$
- 3 $c_2 = 50.0 \div 500.0 = 0.100 \text{ mol dm}^{-3}$

$$0.100 \text{ mol dm}^{-3}$$

Cross-check with the dilution factor: $500.0 \div 25.0 = 20$, and $2.00 \div 20 = 0.100$. ✓ Doing it both ways takes five seconds and catches inverted fractions.

Example 6. A student pipettes 10.0 cm^3 of $0.500 \text{ mol dm}^{-3}$ potassium manganate(VII) into a 100.0 cm^3 volumetric flask and makes it up to the mark. She then pipettes 10.0 cm^3 of *that* solution into a second 100.0 cm^3 flask and makes it up. Calculate the final concentration.

- 1 Each step is a dilution by a factor of $100.0 \div 10.0 = 10$.
- 2 After the first dilution: $0.500 \div 10 = 0.0500 \text{ mol dm}^{-3}$
- 3 After the second dilution: $0.0500 \div 10 = 5.00 \times 10^{-3} \text{ mol dm}^{-3}$

$$5.00 \times 10^{-3} \text{ mol dm}^{-3}$$

Serial dilution factors *multiply*: the overall factor is $10 \times 10 = 100$, not $10 + 10 = 20$. This technique is how the standards for a calibration curve are prepared.

Example 7. In a titration, 25.00 cm³ of sodium hydroxide solution is exactly neutralised by 22.40 cm³ of 0.100 mol dm⁻³ hydrochloric acid. Calculate the concentration of the sodium hydroxide.

- 1 **Equation first:** $\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$ — the ratio is 1 : 1
- 2 $n(\text{HCl}) = cV = 0.100 \times (22.40 \div 1000) = 0.100 \times 0.02240 = 2.240 \times 10^{-3} \text{ mol}$
- 3 From the 1 : 1 ratio, $n(\text{NaOH}) = 2.240 \times 10^{-3} \text{ mol}$
- 4 $c(\text{NaOH}) = n/V = 2.240 \times 10^{-3} \div 0.02500 = 0.0896 \text{ mol dm}^{-3}$

[NaOH] = 0.0896 mol dm⁻³

The three-step shape — known solution → moles → ratio → unknown solution — is the skeleton of every titration question you will ever meet.

Example 8. 25.0 cm³ of sulfuric acid is exactly neutralised by 35.0 cm³ of 0.150 mol dm⁻³ sodium hydroxide. Calculate the concentration of the sulfuric acid.

- 1 $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH(aq)} \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O(l)}$ — the ratio is **1 : 2**, not 1 : 1
- 2 $n(\text{NaOH}) = 0.150 \times 0.0350 = 5.25 \times 10^{-3} \text{ mol}$
- 3 $n(\text{H}_2\text{SO}_4) = 5.25 \times 10^{-3} \div 2 = 2.625 \times 10^{-3} \text{ mol}$
- 4 $c(\text{H}_2\text{SO}_4) = 2.625 \times 10^{-3} \div 0.0250 = 0.105 \text{ mol dm}^{-3}$

[H₂SO₄] = 0.105 mol dm⁻³

Omitting the ÷ 2 gives 0.210 mol dm⁻³ — exactly double, and the single most common wrong answer on diprotic-acid titrations. Write the equation before you write any number.

Example 9. HL STRETCH 50.0 cm³ of 0.200 mol dm⁻³ silver nitrate solution is added to an excess of sodium chloride solution. Calculate the maximum mass of silver chloride precipitate that can form.

- 1 $\text{AgNO}_3(\text{aq}) + \text{NaCl(aq)} \rightarrow \text{AgCl(s)} + \text{NaNO}_3(\text{aq})$, or ionically $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl(s)}$ — ratio 1 : 1
- 2 "An excess of sodium chloride" tells you AgNO_3 is the limiting reactant, so work from it alone.
- 3 $n(\text{AgNO}_3) = 0.200 \times 0.0500 = 0.0100 \text{ mol}$, so $n(\text{AgCl}) = 0.0100 \text{ mol}$
- 4 $M(\text{AgCl}) = 107.87 + 35.45 = 143.32 \text{ g mol}^{-1}$
- 5 $m = 0.0100 \times 143.32 = 1.4332 \text{ g}$

1.43 g

This example chains three sub-topics: concentration (1.4.5) → amount → molar mass (1.4.3) → mass. Recognising "in excess" as the phrase that identifies the limiting reactant saves you an entire calculation.

Sub-topic 1.4.6 — Avogadro's law and gas volumes

SL + HL

Example 10. Nitrogen reacts with hydrogen: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$. Calculate (a) the volume of hydrogen needed to react completely with 200 cm^3 of nitrogen, and (b) the volume of ammonia formed. All volumes are measured at the same temperature and pressure.

- 1 All three species are gases at the same T and p , so by Avogadro's law the volume ratio equals the coefficient ratio, $1 : 3 : 2$.
- 2 (a) $V(\text{H}_2) = 3 \times 200 = 600 \text{ cm}^3$
- 3 (b) $V(\text{NH}_3) = 2 \times 200 = 400 \text{ cm}^3$

(a) 600 cm^3 (b) 400 cm^3

No molar masses, no Avogadro constant, no conversion to mol. When every species is a gas at the same conditions, the coefficients *are* the answer — this is the fastest question type in the whole topic, so do not over-complicate it.

Example 11. 50 cm^3 of methane is mixed with 200 cm^3 of oxygen and ignited. After the reaction the mixture is cooled to the original temperature and pressure. Calculate the total volume of gas remaining.

- 1 $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ — note the state symbol after cooling
- 2 **Limiting reactant test:** CH_4 : $50 \div 1 = 50$. O_2 : $200 \div 2 = 100$. The smaller value wins, so **methane is limiting**.
- 3 O_2 used = $2 \times 50 = 100 \text{ cm}^3$, so O_2 remaining = $200 - 100 = 100 \text{ cm}^3$
- 4 CO_2 formed = $1 \times 50 = 50 \text{ cm}^3$
- 5 Water has condensed to a liquid on cooling, so it contributes no volume.

Total = $50 \text{ cm}^3 \text{ CO}_2$ + 100 cm^3 excess O_2 = 150 cm^3

Two marks are routinely dropped here: forgetting the leftover oxygen, and counting the water. Read the temperature statement before you total anything.

Example 12. 20 cm^3 of propane is burned in 150 cm^3 of oxygen. Determine the total volume of gas present (a) at 120°C and (b) after cooling to 20°C . All volumes are measured at the same pressure.

- 1 $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}$
- 2 **Limiting:** C_3H_8 : $20 \div 1 = 20$. O_2 : $150 \div 5 = 30$. Propane is limiting.
- 3 O_2 used = $5 \times 20 = 100 \text{ cm}^3 \rightarrow$ excess $\text{O}_2 = 50 \text{ cm}^3$
- 4 $\text{CO}_2 = 3 \times 20 = 60 \text{ cm}^3$; $\text{H}_2\text{O} = 4 \times 20 = 80 \text{ cm}^3$
- 5 (a) **At 120°C** water is a *gas* and counts: $60 + 80 + 50 = 190 \text{ cm}^3$
- 6 (b) **At 20°C** water has condensed: $60 + 50 = 110 \text{ cm}^3$

(a) 190 cm^3 (b) 110 cm^3

The 80 cm^3 difference is entirely the water. Examiners set this pair deliberately to test whether you read the temperature — treat any stated temperature above 100°C as an instruction to include the water.

Example 13. Calculate the volume occupied at STP by (a) 0.250 mol of carbon dioxide and (b) 0.250 mol of helium.

- 1 At STP, $V = n \times V_m = n \times 22.7 \text{ dm}^3 \text{ mol}^{-1}$
- 2 (a) $V = 0.250 \times 22.7 = 5.675 \text{ dm}^3$
- 3 (b) $V = 0.250 \times 22.7 = 5.675 \text{ dm}^3$ — identical

(a) 5.68 dm³ (b) 5.68 dm³

This is Avogadro's law in action. Carbon dioxide molecules are eleven times more massive than helium atoms, but the volumes are the same because volume depends only on the number of particles. If a question asks you to "explain why the two volumes are equal", that sentence is the answer.

Example 14. Calculate the volume occupied by 4.40 g of carbon dioxide at STP.

- 1 Route: mass → amount → volume. Never jump straight from mass to volume.
- 2 $M(\text{CO}_2) = 12.01 + 2(16.00) = 44.01 \text{ g mol}^{-1}$
- 3 $n = 4.40 \div 44.01 = 0.09998 \text{ mol}$
- 4 $V = 0.09998 \times 22.7 = 2.2695 \text{ dm}^3$

2.27 dm³ (equivalently 2270 cm³)

Sanity check: 4.40 g is almost exactly one-tenth of a mole of CO₂, so the volume should be almost exactly one-tenth of 22.7. It is.

Example 15. HL STRETCH 25.0 cm³ of 2.00 mol dm⁻³ hydrochloric acid is added to an excess of calcium carbonate chips. Calculate the volume of carbon dioxide produced, measured at STP.

- 1 $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
- 2 Carbonate is in excess, so the acid is limiting: $n(\text{HCl}) = 2.00 \times 0.0250 = 0.0500 \text{ mol}$
- 3 From the 2 : 1 ratio, $n(\text{CO}_2) = 0.0500 \div 2 = 0.0250 \text{ mol}$
- 4 $V = 0.0250 \times 22.7 = 0.5675 \text{ dm}^3$

0.568 dm³, or 568 cm³

This is the classic cross-over question: it starts in solution (1.4.5), passes through a mole ratio, and finishes as a gas volume (1.4.6 / 1.5.3). Note that the calcium carbonate is a *solid*, so Avogadro's law says nothing about it — only the CO₂ has a gas volume.

Example 16. **HL STRETCH** A pure gaseous element has a density of 1.25 g dm^{-3} at STP. (a) Determine its molar mass. (b) Suggest the identity of the gas, giving a reason.

- 1 Consider 1.00 dm^3 of the gas. Its mass is 1.25 g .
- 2 n in $1.00 \text{ dm}^3 = 1.00 \div 22.7 = 0.04405 \text{ mol}$
- 3 $M = m/n = 1.25 \div 0.04405 = 28.4 \text{ g mol}^{-1}$ (equivalently $M = \text{density} \times 22.7 = 1.25 \times 22.7 = 28.4$)
- 4 (b) An *element* with $M \approx 28.4 \text{ g mol}^{-1}$: nitrogen exists as N_2 , for which $M = 2 \times 14.01 = 28.02 \text{ g mol}^{-1}$. Silicon has $A_r = 28.09$ but is a solid at STP.

(a) 28.4 g mol^{-1} (b) Nitrogen, N_2

The question says "element", which rules out carbon monoxide ($M = 28.01$ — an equally good numerical fit). Read the stem. Note also that $M = \text{density} \times V_m$ works for any gas at any conditions, provided you use the molar volume for *those* conditions.

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PART E — PRACTICE

SECTION E1

Paper 1A style — multiple choice

SL + HL

20 questions. Target time: 30 minutes. Choose the one best answer. A calculator and the data booklet are permitted, but train yourself to eliminate options by estimation first. Take $V_m = 22.7 \text{ dm}^3 \text{ mol}^{-1}$ at STP. Mark scheme with full explanations begins after Section E3.

1. What is the concentration of a solution containing 0.20 mol of solute in 500 cm^3 of solution?
A. 0.10 mol dm^{-3}
B. 0.40 mol dm^{-3}
C. 4.0 mol dm^{-3}
D. 100 mol dm^{-3}
2. What mass of sodium hydroxide is needed to prepare 250 cm^3 of a $0.200 \text{ mol dm}^{-3}$ solution?
A. 1.00 g
B. 2.00 g
C. 8.00 g
D. 20.0 g
3. What is the concentration of chloride ions in a 0.10 mol dm^{-3} solution of magnesium chloride, MgCl_2 ?
A. $0.050 \text{ mol dm}^{-3}$
B. 0.10 mol dm^{-3}
C. 0.20 mol dm^{-3}
D. 0.30 mol dm^{-3}
4. 25.0 cm^3 of $0.400 \text{ mol dm}^{-3}$ sodium chloride solution is diluted to a total volume of 100.0 cm^3 . What is the concentration of the diluted solution?
A. $0.0250 \text{ mol dm}^{-3}$
B. $0.100 \text{ mol dm}^{-3}$
C. $0.160 \text{ mol dm}^{-3}$
D. 1.60 mol dm^{-3}
5. A solution of copper(II) sulfate has a concentration of $0.0200 \text{ mol dm}^{-3}$. What is this concentration in g dm^{-3} ? ($M_r(\text{CuSO}_4) = 159.62$)
A. 0.319 g dm^{-3}
B. 3.19 g dm^{-3}
C. 8.00 g dm^{-3}
D. 31.9 g dm^{-3}

6. Which statement is a correct expression of Avogadro's law?

- A. One mole of any gas contains 6.02×10^{23} atoms.
- B. Equal masses of all gases at the same temperature and pressure occupy equal volumes.
- C. Equal volumes of all gases at the same temperature and pressure contain equal numbers of molecules.
- D. The volume of a gas is inversely proportional to the amount of gas at constant temperature and pressure.

7. 60 cm^3 of nitrogen reacts with an excess of hydrogen: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$. What volume of ammonia is formed, measured at the same temperature and pressure?

- A. 30 cm^3
- B. 60 cm^3
- C. 120 cm^3
- D. 180 cm^3

8. What volume is occupied by 0.100 mol of an ideal gas at STP?

- A. 1.14 dm^3
- B. 2.27 dm^3
- C. 22.7 dm^3
- D. 227 dm^3

9. Four separate 1.00 dm^3 containers hold H_2 , O_2 , CO_2 and NH_3 , all at the same temperature and pressure. Which container holds the greatest number of molecules?

- A. H_2
- B. O_2
- C. CO_2
- D. They all contain the same number of molecules.

10. What mass of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, is present in 500 cm^3 of a $0.100 \text{ mol dm}^{-3}$ solution?

- A. 0.901 g
- B. 9.01 g
- C. 18.0 g
- D. 90.1 g

11. 20.0 cm^3 of sodium hydroxide solution is exactly neutralised by 25.0 cm^3 of $0.100 \text{ mol dm}^{-3}$ hydrochloric acid. What is the concentration of the sodium hydroxide?

- A. $0.0800 \text{ mol dm}^{-3}$
- B. $0.100 \text{ mol dm}^{-3}$
- C. $0.125 \text{ mol dm}^{-3}$
- D. $0.250 \text{ mol dm}^{-3}$

12. What volume of $0.100 \text{ mol dm}^{-3}$ sodium hydroxide is needed to neutralise exactly 10.0 cm^3 of $0.200 \text{ mol dm}^{-3}$ sulfuric acid?

- A. 10.0 cm^3
- B. 20.0 cm^3
- C. 40.0 cm^3
- D. 80.0 cm^3

13. Which 1.0 g sample occupies the greatest volume at STP?

- A. H_2
- B. He
- C. CH_4
- D. CO_2

14. 100 cm^3 of carbon monoxide is mixed with 100 cm^3 of oxygen and ignited: $2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g})$. What is the total volume of gas after the reaction, measured at the same temperature and pressure?

- A. 100 cm^3
- B. 150 cm^3
- C. 200 cm^3
- D. 250 cm^3

15. 100 cm^3 of $0.200 \text{ mol dm}^{-3}$ sodium chloride solution is mixed with 100 cm^3 of distilled water. What is the concentration of the resulting solution?

- A. $0.0500 \text{ mol dm}^{-3}$
- B. $0.100 \text{ mol dm}^{-3}$
- C. $0.200 \text{ mol dm}^{-3}$
- D. $0.400 \text{ mol dm}^{-3}$

16. 10.0 cm^3 of $0.500 \text{ mol dm}^{-3}$ sodium hydroxide is added to 10.0 cm^3 of $0.500 \text{ mol dm}^{-3}$ sodium hydroxide. What is the concentration of the resulting solution?

- A. $0.250 \text{ mol dm}^{-3}$
- B. $0.500 \text{ mol dm}^{-3}$
- C. 1.00 mol dm^{-3}
- D. 2.00 mol dm^{-3}

17. What amount of gas, in mol, is present in 0.500 dm^3 of an ideal gas at STP?

- A. 0.0220 mol
- B. 0.0446 mol
- C. 11.4 mol
- D. 45.4 mol

18. A gas has a density of 1.42 g dm^{-3} at STP. What is its molar mass?

- A. 16.0 g mol^{-1}
- B. 32.2 g mol^{-1}
- C. 44.0 g mol^{-1}
- D. 64.0 g mol^{-1}

19. 25 cm^3 of propane is burned in 200 cm^3 of oxygen: $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l})$. What is the total volume of gas remaining after the mixture is cooled to the original room temperature and pressure?

- A. 75 cm^3
- B. 100 cm^3
- C. 150 cm^3
- D. 250 cm^3

20. Which statement is correct?

- A. Square brackets, as in $[\text{HCl}]$, denote a concentration in g dm^{-3} .
- B. Avogadro's law can be applied to a solid provided the temperature and pressure are stated.
- C. The volume V in the relationship $c = n/V$ is the volume of solvent added.
- D. Diluting a solution decreases its concentration but does not change the amount of solute present.

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SECTION E2

Paper 1B style — data-based questions

SL + HL

2 questions, 20 marks. Target time: 30 minutes. Calculator and data booklet permitted. Show all working — method marks are available even where the final answer is wrong.

Question 1. A student prepares a standard solution of anhydrous sodium carbonate and then dilutes it. [10 marks]
The following data are recorded.

MEASUREMENT	VALUE
Mass of Na_2CO_3 weighed out	2.650 g
Volume of volumetric flask used (solution A)	250.0 cm^3
Volume of solution A pipetted into a second flask	25.00 cm^3
Volume of second volumetric flask (solution B)	250.0 cm^3

(a) State what is meant by a *standard solution*. [1]

(b) Calculate the amount, in mol, of sodium carbonate weighed out, and hence the concentration of solution A in mol dm^{-3} . [3]

(c) Determine the concentration of solution B. [2]

(d) Deduce the concentration of sodium ions, $[\text{Na}^+]$, in solution B. [1]

(e) Explain why a volumetric flask, rather than a measuring cylinder, is used to make up solution A to 250.0 cm^3 . [1]

(f) A second student adds distilled water past the graduation mark of the 250.0 cm^3 flask. State and explain the effect this has on the calculated concentration of solution A compared with its true value. [2]

Question 2. A student determines the molecular formula of a gaseous hydrocarbon, C_xH_y . 10 cm^3 of the hydrocarbon is mixed with 100 cm^3 of oxygen (an excess) in a eudiometer and sparked. The following volumes are measured, all at the same temperature and pressure.

STAGE	VOLUME OF GAS / CM^3
Hydrocarbon before reaction	10
Oxygen before reaction	100
After sparking and cooling to the original temperature	80
After passing the remaining gas through aqueous sodium hydroxide	50

(a) State Avogadro's law. [1]

(b) Aqueous sodium hydroxide absorbs carbon dioxide. Deduce the volume of carbon dioxide produced. [1]

(c) Explain why the mixture was cooled to the original temperature before the volume was measured. [1]

(d) Determine the volume of oxygen that reacted. [2]

(e) Deduce the value of x , explaining your reasoning. [2]

(f) Using the general equation $C_xH_y + (x + y/4)O_2 \rightarrow xCO_2 + (y/2)H_2O$, deduce the value of y . [2]

(g) State the molecular formula and name of the hydrocarbon. [1]

SECTION E3

Paper 2 style — structured questions

SL + HL

3 questions, 30 marks (plus 3 HL Stretch marks). Target time: 50 minutes. Calculator and data booklet permitted.

Question 1. The concentration of ethanoic acid, CH_3COOH , in a sample of vinegar is determined by [10 marks] titration. A 25.00 cm^3 portion of the vinegar is transferred to a 250.0 cm^3 volumetric flask and made up to the mark with distilled water. A 25.00 cm^3 portion of this diluted solution then requires 21.50 cm^3 of $0.1000 \text{ mol dm}^{-3}$ sodium hydroxide for exact neutralisation.

- (a) Write the balanced equation for the reaction between ethanoic acid and sodium hydroxide. [1]
- (b) Calculate the amount, in mol, of sodium hydroxide used. [1]
- (c) Determine the concentration of ethanoic acid in the *diluted* solution. [2]
- (d) Determine the concentration of ethanoic acid in the *original* vinegar, in mol dm^{-3} . [2]
- (e) Vinegar is sold with its acid content labelled in g per 100 cm^3 . Calculate this value for the sample. [3]
- (f) Suggest one reason why the vinegar was diluted before being titrated. [1]

Question 2. 0.240 g of magnesium ribbon is added to 50.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid. The [10 marks] hydrogen produced is collected and its volume measured at STP.

- (a) Write the balanced equation, including state symbols, for the reaction. [1]
- (b) Calculate the amount, in mol, of magnesium and the amount, in mol, of hydrochloric acid. [2]
- (c) Deduce, with justification, which reactant is limiting. [2]
- (d) Determine the volume of hydrogen produced at STP, in cm^3 . [3]
- (e) Assuming the volume of the solution does not change, calculate the concentration of hydrochloric acid remaining after the reaction is complete. [2]

Question 3. Ammonia is manufactured by the Haber process: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$. [10 marks + 3 HL Stretch]

In a laboratory simulation, 60.0 dm^3 of nitrogen is passed over the catalyst with an excess of hydrogen. All gas volumes are measured at STP.

- (a) Using Avogadro's law, deduce the volume of hydrogen that would react completely with 60.0 dm^3 of nitrogen, and the maximum volume of ammonia that could be formed. [2]
- (b) In practice only 40.0 dm^3 of ammonia is obtained. Calculate the percentage yield. [2]
- (c) Determine the amount, in mol, of ammonia in 40.0 dm^3 measured at STP. [2]
- (d) All of this ammonia is dissolved in water and the solution made up to 2.00 dm^3 . Calculate the concentration of the resulting solution in (i) mol dm^{-3} and (ii) g dm^{-3} . [3]
- (e) Explain why the total volume of gas decreases as the reaction proceeds. [1]

HL STRETCH EXTENSION PART (F) — BACK TITRATION

(f) A *different* batch of aqueous ammonia is analysed as follows. A 25.00 cm^3 portion is added to 50.00 cm^3 of $0.500 \text{ mol dm}^{-3}$ hydrochloric acid, which is in excess. The unreacted acid then requires 18.60 cm^3 of $0.400 \text{ mol dm}^{-3}$ sodium hydroxide for neutralisation. Determine the concentration of ammonia in this batch. [3]

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PART F — MARK SCHEMES

MARK SCHEME E1

Paper 1A style — answer key

SL + HL

Q	1	2	3	4	5	6	7	8	9	10
Ans	B	B	C	B	B	C	C	B	D	B
Q	11	12	13	14	15	16	17	18	19	20
Ans	C	C	A	B	B	B	A	B	C	D

EXPLANATIONS — INCLUDING WHY THE DISTRACTORS ARE WRONG

1 — B. $V = 500 \div 1000 = 0.500 \text{ dm}^3$, so $c = 0.20 \div 0.500 = 0.40 \text{ mol dm}^{-3}$. Option D (100) comes from multiplying by 500 instead of dividing by 0.500; option A halves rather than doubles.

2 — B. $n = cV = 0.200 \times 0.250 = 0.0500 \text{ mol}$; $M(\text{NaOH}) = 22.99 + 16.00 + 1.01 = 40.00 \text{ g mol}^{-1}$; $m = 0.0500 \times 40.00 = 2.00 \text{ g}$. Option C (8.00 g) is the mass needed for 1 dm^3 — a failure to scale for the 250 cm^3 flask.

3 — C. $\text{MgCl}_2 \rightarrow \text{Mg}^{2+} + 2\text{Cl}^-$, so $[\text{Cl}^-] = 2 \times 0.10 = 0.20 \text{ mol dm}^{-3}$. Option B is the trap: it quotes the concentration of the compound rather than of the ion asked for.

4 — B. $c_1V_1 = c_2V_2$: $0.400 \times 25.0 = c_2 \times 100.0$, so $c_2 = 0.100 \text{ mol dm}^{-3}$. The dilution factor is 4, and $0.400 \div 4 = 0.100$. Option D inverts the ratio and multiplies by 4 instead.

5 — B. $\text{g dm}^{-3} = \text{mol dm}^{-3} \times M = 0.0200 \times 159.62 = 3.19 \text{ g dm}^{-3}$. Option A is a decimal-place slip; option D is ten times too large.

6 — C. This is the syllabus statement verbatim. A fails for any molecular gas (1 mol of O_2 contains 1.2×10^{24} atoms, not 6.02×10^{23}). B is wrong because equal *masses* correspond to different amounts for gases of different molar mass. D reverses the relationship — volume is *directly* proportional to amount.

7 — C. All species are gases at the same T and p , so the volume ratio is the coefficient ratio $\text{N}_2 : \text{NH}_3 = 1 : 2$. $V(\text{NH}_3) = 2 \times 60 = 120 \text{ cm}^3$. Option D (180 cm^3) is the volume of *hydrogen* required — read which gas is asked for.

8 — B. $V = n \times V_m = 0.100 \times 22.7 = 2.27 \text{ dm}^3$. Option C is the molar volume itself, which corresponds to 1.00 mol .

9 — D. This is Avogadro's law stated as a question. Equal volumes of any gases at the same temperature and pressure contain equal numbers of molecules, so all four containers hold the same number. The identity, mass and size of the molecules are all irrelevant. Choosing A or C means you have reasoned about mass or molecular size rather than about particle count.

10 — B. $n = 0.100 \times 0.500 = 0.0500 \text{ mol}$; $M(\text{C}_6\text{H}_{12}\text{O}_6) = 180.18 \text{ g mol}^{-1}$; $m = 0.0500 \times 180.18 = 9.01 \text{ g}$. Option D is the mass in 1 dm^3 ; option A is a factor-of-ten slip.

11 – C. The reaction is 1 : 1. $n(\text{HCl}) = 0.100 \times 0.0250 = 2.50 \times 10^{-3} \text{ mol} = n(\text{NaOH})$. $c(\text{NaOH}) = 2.50 \times 10^{-3} \div 0.0200 = 0.125 \text{ mol dm}^{-3}$. Option A (0.0800) comes from dividing the volumes the wrong way round. Note the sense-check: a *smaller* volume of base neutralised a larger volume of acid, so the base must be *more* concentrated than 0.100.

12 – C. $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$. $n(\text{H}_2\text{SO}_4) = 0.200 \times 0.0100 = 2.00 \times 10^{-3} \text{ mol}$, so $n(\text{NaOH}) = 4.00 \times 10^{-3} \text{ mol}$ and $V = 4.00 \times 10^{-3} \div 0.100 = 0.0400 \text{ dm}^3 = 40.0 \text{ cm}^3$. Option B (20.0 cm³) omits the factor of 2 for the diprotic acid.

13 – A. The greatest volume belongs to the greatest *amount*, so convert 1.0 g of each to mol: $\text{H}_2 \rightarrow 1.0 \div 2.02 = 0.50 \text{ mol}$; $\text{He} \rightarrow 0.25 \text{ mol}$; $\text{CH}_4 \rightarrow 0.062 \text{ mol}$; $\text{CO}_2 \rightarrow 0.023 \text{ mol}$. The lightest gas gives the most particles per gram and therefore the largest volume. Choosing D confuses molar mass with volume.

14 – B. Limiting-reactant test: $\text{CO } 100 \div 2 = 50$; $\text{O}_2 100 \div 1 = 100$, so CO is limiting. O_2 used = 50 cm³, leaving 50 cm³ unreacted; CO_2 formed = 100 cm³. Total = 100 + 50 = 150 cm³. Option C (200) forgets that only half the oxygen reacts and counts the leftover incorrectly; option A ignores the excess oxygen entirely.

15 – B. The total volume doubles from 100 cm³ to 200 cm³ while the amount of solute is unchanged, so the concentration halves: $0.200 \div 2 = 0.100 \text{ mol dm}^{-3}$.

16 – B. Concentration is an **intensive** property: it does not depend on how much of the solution you have. Combining two portions of identical solution doubles both the amount of solute and the volume, so the ratio is unchanged at 0.500 mol dm⁻³. Option C (1.00) treats concentrations as though they add, which is the single most revealing misconception in this sub-topic.

17 – A. $n = V \div V_m = 0.500 \div 22.7 = 0.0220 \text{ mol}$. Option C (11.4) multiplies instead of dividing; option B uses $22.4 \div 2$ rather than the IB value of 22.7.

18 – B. $M = \text{density} \times V_m = 1.42 \times 22.7 = 32.2 \text{ g mol}^{-1}$, consistent with oxygen, O_2 ($M = 32.00$). Option C would require a density of 1.94 g dm⁻³.

19 – C. Limiting test: $\text{C}_3\text{H}_8 25 \div 1 = 25$; $\text{O}_2 200 \div 5 = 40$, so propane is limiting. O_2 used = $5 \times 25 = 125 \text{ cm}^3$, leaving 75 cm³; CO_2 formed = $3 \times 25 = 75 \text{ cm}^3$. The water is a liquid at room temperature and contributes nothing. Total = 75 + 75 = 150 cm³. Option A counts only the CO_2 and forgets the excess oxygen.

20 – D. Dilution adds solvent, increasing V while n stays constant, so $c = n/V$ falls. A is wrong because square brackets denote *molar* concentration; B is wrong because Avogadro's law applies only to gases; C is wrong because V is the volume of the final *solution*, not of the solvent added.

MARK SCHEME E2

Paper 1B style

SL + HL

"ECF" = error carried forward: a correct method applied to an earlier incorrect answer still earns the later marks. Accept answers within reasonable rounding. "OWTTE" = or words to that effect.

Question 1 — Standard solution and dilution [10 marks]

PART	EXPECTED ANSWER	MARKS
(a)	A solution of accurately known concentration [1] <i>Accept "prepared by dissolving a known mass of solute and making up to a known volume" OWTTE.</i>	1
(b)	$M(\text{Na}_2\text{CO}_3) = 2(22.99) + 12.01 + 3(16.00) = 105.99 \text{ g mol}^{-1}$ [1] $n = 2.650 \div 105.99 = 0.02500 \text{ mol}$ [1] $c = 0.02500 \div 0.2500 = 0.1000 \text{ mol dm}^{-3}$ [1] <i>Award the third mark only if the volume has been converted to dm³. Unit required.</i>	3
(c)	Dilution factor = $250.0 \div 25.00 = 10$ [1] $c(\text{B}) = 0.1000 \div 10 = 0.01000 \text{ mol dm}^{-3}$ [1] <i>Accept the equivalent route using $c_1V_1 = c_2V_2$. ECF from (b).</i>	2
(d)	$\text{Na}_2\text{CO}_3 \rightarrow 2\text{Na}^+ + \text{CO}_3^{2-}$, so $[\text{Na}^+] = 2 \times 0.01000 = 0.0200 \text{ mol dm}^{-3}$ [1] (ECF)	1
(e)	A volumetric flask is calibrated to deliver one precise total volume and has a far smaller uncertainty than a measuring cylinder, so the concentration is known accurately [1] OWTTE	1
(f)	The actual volume of solution is greater than 250.0 cm^3 , so the <i>true</i> concentration is lower than $0.1000 \text{ mol dm}^{-3}$ [1] The calculation assumes exactly 250.0 cm^3 , so the <i>calculated</i> value is higher than the true value [1] OWTTE <i>Both marks require a clear statement of direction. Do not credit "it would be inaccurate" alone.</i>	2

Question 2 — Gas volumes and molecular formula [10 marks]

PART	EXPECTED ANSWER	MARKS
(a)	Equal volumes of all gases measured at the same temperature and pressure contain equal numbers of molecules [1] <i>Accept "$V \propto n$ at constant T and p". Do not award if the T and p condition is omitted.</i>	1
(b)	$V(\text{CO}_2) = 80 - 50 = 30 \text{ cm}^3$ [1]	1
(c)	So that the water formed condenses to a liquid and does not contribute to the gas volume / so that all volumes are compared at the same temperature and pressure, as Avogadro's law requires [1] OWTTE	1
(d)	Gas remaining after removing $\text{CO}_2 = 50 \text{ cm}^3$, which is the unreacted oxygen [1] $V(\text{O}_2 \text{ reacted}) = 100 - 50 = 50 \text{ cm}^3$ [1]	2
(e)	By Avogadro's law the volume ratio equals the mole ratio: $\text{C}_x\text{H}_y : \text{CO}_2 = 10 : 30 = 1 : 3$ [1] Each molecule therefore contains 3 carbon atoms, so $x = 3$ [1]	2
(f)	$\text{C}_x\text{H}_y : \text{O}_2 = 10 : 50 = 1 : 5$, so $x + y/4 = 5$ [1] $3 + y/4 = 5 \rightarrow y/4 = 2 \rightarrow y = 8$ [1] (ECF from (d) and (e))	2
(g)	C_3H_8, propane [1] <i>Both the formula and the name are needed for the mark.</i>	1

Verification: $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$. 10 cm^3 of propane consumes 50 cm^3 of oxygen and produces 30 cm^3 of CO_2 . After cooling: 30 (CO_2) + 50 (excess O_2) = 80 cm^3 . ✓ After absorbing CO_2 : 50 cm^3 . ✓ Every figure in the table is accounted for.

MARK SCHEME E3

Paper 2 style

SL + HL

Question 1 — Vinegar titration [10 marks]

PART	EXPECTED ANSWER	MARKS
(a)	$\text{CH}_3\text{COOH}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{CH}_3\text{COONa}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ [1] <i>Accept correct formulae without state symbols. The ratio is 1 : 1 — ethanoic acid is monoprotic.</i>	1
(b)	$n(\text{NaOH}) = cV = 0.1000 \times 0.02150 = 2.150 \times 10^{-3} \text{ mol}$ [1]	1
(c)	1 : 1 ratio, so $n(\text{CH}_3\text{COOH}) = 2.150 \times 10^{-3} \text{ mol}$ [1] $c = 2.150 \times 10^{-3} \div 0.02500 = 0.08600 \text{ mol dm}^{-3}$ [1]	2
(d)	Dilution factor = $250.0 \div 25.00 = 10$ [1] $c(\text{vinegar}) = 0.08600 \times 10 = 0.8600 \text{ mol dm}^{-3}$ [1] (ECF) <i>Multiplying rather than dividing is essential here — the original is the more concentrated solution.</i>	2
(e)	$M(\text{CH}_3\text{COOH}) = 2(12.01) + 4(1.01) + 2(16.00) = 60.06 \text{ g mol}^{-1}$ [1] Mass concentration = $0.8600 \times 60.06 = 51.65 \text{ g dm}^{-3}$ [1] Per $100 \text{ cm}^3 = 51.65 \div 10 = 5.17 \text{ g}$ [1] <i>Accept 5.2 g. A typical table vinegar is 4–6 %, so this is a sensible result.</i>	3
(f)	Undiluted vinegar would require a titre far larger than the capacity of a 50 cm^3 burette / repeated refilling would increase the uncertainty [1] OWTTE <i>Accept "to bring the titre into the accurate working range of the burette".</i>	1

Question 2 — Magnesium and hydrochloric acid [10 marks]

PART	EXPECTED ANSWER	MARKS
(a)	$\text{Mg}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})$ [1] <i>State symbols required as the question asks for them.</i>	1
(b)	$n(\text{Mg}) = 0.240 \div 24.31 = 9.87 \times 10^{-3} \text{ mol}$ [1] $n(\text{HCl}) = 1.00 \times 0.0500 = 0.0500 \text{ mol}$ [1]	2
(c)	HCl required to react with all the Mg = $2 \times 9.87 \times 10^{-3} = 0.0197 \text{ mol}$ [1] 0.0500 mol of HCl is available, which is more than 0.0197 mol, so the acid is in excess and magnesium is limiting [1] <i>Award the second mark only if the comparison is explicit. A bare "Mg is limiting" scores 0.</i>	2
(d)	Mg : H ₂ is 1 : 1, so $n(\text{H}_2) = 9.87 \times 10^{-3} \text{ mol}$ [1] $V = n \times 22.7 = 9.87 \times 10^{-3} \times 22.7 = 0.2241 \text{ dm}^3$ [1] = 224 cm³ [1] <i>The final mark is for the conversion to cm³ as the question requires. Accept 224–225 cm³.</i>	3
(e)	$n(\text{HCl}) \text{ remaining} = 0.0500 - 0.0197 = 0.0303 \text{ mol}$ [1] $c = 0.0303 \div 0.0500 = 0.605 \text{ mol dm}^{-3}$ [1] (ECF) <i>Accept 0.61 mol dm⁻³.</i>	2

Question 3 — The Haber process [10 marks + 3 HL Stretch]

PART	EXPECTED ANSWER	MARKS
(a)	Volume ratio = coefficient ratio $N_2 : H_2 : NH_3 = 1 : 3 : 2$ $V(H_2) = 3 \times 60.0 = 180 \text{ dm}^3$ [1] $V(NH_3) = 2 \times 60.0 = 120 \text{ dm}^3$ [1]	2
(b)	% yield = (actual $\div$ theoretical) $\times 100 = (40.0 \div 120) \times 100$ [1] = 33.3 % [1] (ECF from (a))	2
(c)	$n = V \div V_m = 40.0 \div 22.7$ [1] = 1.76 mol [1]	2
(d)	(i) $c = 1.762 \div 2.00 = 0.881 \text{ mol dm}^{-3}$ [2] (ii) $M(NH_3) = 14.01 + 3(1.01) = 17.04 \text{ g mol}^{-1}$; $0.881 \times 17.04 = 15.0 \text{ g dm}^{-3}$ [1] ECF throughout. Accept 0.88 and 15 g dm ⁻³ .	3
(e)	Four moles of gas (1 $N_2 + 3 H_2$) produce only two moles of gas (2 NH_3); since volume is proportional to amount of gas at constant T and p , the total volume falls [1] OWTTE	1
(f) HL	$n(HCl)$ added = $0.500 \times 0.05000 = 0.02500 \text{ mol}$; $n(NaOH) = 0.400 \times 0.01860 = 7.440 \times 10^{-3} \text{ mol}$, which equals the unreacted HCl [1] $n(HCl)$ that reacted with the ammonia = $0.02500 - 0.007440 = 0.01756 \text{ mol}$ [1] $NH_3 + HCl \rightarrow NH_4Cl$ is 1 : 1, so $c(NH_3) = 0.01756 \div 0.02500 = 0.702 \text{ mol dm}^{-3}$ [1] The structure of every back titration: total moles of reagent added, minus moles found in excess, equals moles that reacted with the unknown.	3

SELF-ASSESSMENT: WHAT YOUR SCORE IS TELLING YOU

SECTION E1 (MCQ)	DIAGNOSIS AND NEXT STEP
17–20	Secure. Structure 1.4 is complete — move on to Structure 1.5 (ideal gases, $pV = nRT$ and deviations from ideality) and then to Reactivity 2.1.
13–16	The chemistry is sound but unit conversion or the limiting-reactant test is slipping. Re-do Solved Examples 5, 11 and 12 from a blank page.
9–12	Re-read Part A in full and rebuild the solution–gas road map from memory before attempting Sections E2 and E3.
0–8	Go back to first principles: practise converting ten volumes from cm ³ to dm ³ , then work $n = cV$ in both directions on ten solutions of your own devising before touching exam questions.



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