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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Reactivity 2.1
Equations and Reacting Quantities

Booklet 1 · Understandings 2.1.1 – 2.1.2

Balanced equations with state symbols, and how the mole ratio turns them into reacting masses, gas volumes and concentrations, including titrations and back titrations

Complete revision notes, 18 worked examples, model answers & 35 practice questions · Standard Level & Higher Level

How the level tags work in this booklet

SL + HL Every understanding in Reactivity 2.1 is common to both levels, so the whole booklet is core material.

HL Reactivity 2.1 has no HL-only content. The ideal gas equation (Structure 1.5) and redox titrations (Reactivity 3.2) are SL + HL skills from other topics and are labelled where they appear. ★ marks harder, multi-step examples and questions.

Inside this booklet

1. What the syllabus actually asks of you
2. 2.1.1 — Chemical equations
3. 2.1.2 — The mole toolkit
4. 2.1.2 — Reacting masses
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6. 2.1.2 — Solutions and titration
7. 2.1.2 — Back titration
8. Key points to memorise
9. Check yourself
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1 What the syllabus actually asks of you

Reactivity 2 asks three questions about any reaction: how much, how fast and how far? Reactivity 2.1 answers the first. It is the toolkit for nearly every calculation in the course: you write a balanced equation, read the mole ratio from it, and use that ratio to find the masses, gas volumes and concentrations of reactants and products. This is **stoichiometry**, from the Greek *stoicheion* (element) and *metron* (measure). Expect it in every paper: short calculations in Paper 1A, data-based titration questions in Paper 1B, and multi-step calculations inside larger questions in Paper 2.

Understanding	What it says	What you must be able to do	Level
2.1.1 Chemical equations	Chemical equations show the ratio of reactants and products in a reaction.	Deduce chemical equations when reactants and products are specified. **Include the use of state symbols in chemical equations.**	SL+HL
2.1.2 Mole ratios	The mole ratio of an equation can be used to determine the masses and/or volumes of reactants and products, and the concentrations of reactants and products for reactions occurring in solution.	Calculate reacting masses and/or volumes and concentrations of reactants and products. Avogadro's law and the definitions of molar concentration are covered in Structure 1.4. **The values for A_r given in the data booklet to two decimal places should be used in calculations.**	SL+HL

Understandings 2.1.3 (limiting reactant and theoretical yield), 2.1.4 (percentage yield) and 2.1.5 (atom economy) are covered in Booklet 2.

Linking questions and connections for 2.1.1 – 2.1.2

- **Reactivity 3.2:** when is it useful to use half-equations? → section 2.5
- **Structure 1.5:** how does the molar volume of a gas vary with changes in temperature and pressure? → section 5.4
- **Nature of Science, Structure 1.4:** in what ways does Avogadro's law help us to describe, but not explain, the behaviour of gases? → section 5.1
- **Structure 1.4:** the mole, molar mass, molar concentration and molar volume are the tools used throughout → section 3
- **Reactivity 3.1 and 3.2:** acid–base and redox titrations → sections 6 and 7
- **Tool 1 and Inquiry:** burette readings, uncertainty and concordant titres → section 6.3
- **Reactivity 1.3:** combustion equations and combustion analysis use the same balancing skills.

What the data booklet does and does not give you

It gives you: relative atomic masses to two decimal places (use exactly these: C = 12.01, H = 1.01, O = 16.00); the Avogadro constant, 6.02×10^{23} mol^{-1}; the molar volume of an ideal gas at STP, 22.7 dm^3 mol^{-1}, and the STP conditions (273 K, 100 kPa); the gas constant, $R = 8.31$ J K^{-1} mol^{-1}; and key formulas, including the ideal gas equation.

It does not give you: the balanced equation or any mole ratio; formulas of ions and compounds; state symbols; unit conversions (cm^3 → dm^3, °C → K, kPa → Pa); or how to treat burette readings and titres. All of these must be learned.

2 Reactivity 2.1.1 — Chemical equations SL + HL

2.1 What an equation tells you

In 1789 Antoine Lavoisier showed by careful weighing that matter keeps its mass through a chemical change: this is the **law of conservation of mass**. His quantitative work also overturned the "phlogiston" theory of burning, which could not explain why metals *gain* mass when they burn. A chemical equation is the book-keeping that follows from this law. Atoms are rearranged but never created or destroyed, so every element must have the same number of atoms on each side.

> **Definitions to learn**
>
> A **chemical equation** uses chemical symbols to show the reactants (left) and products (right) of a reaction, in the simplest whole-number ratio of the particles that react.
>
> **Stoichiometric coefficients** are the numbers in front of each formula. In a **balanced** equation they make the number of atoms of each element (and the total charge) the same on both sides. They give the **mole ratio** of the reaction.

For $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, the coefficients say that 2 molecules of hydrogen react with 1 molecule of oxygen to give 2 molecules of water. Multiplying by the Avogadro constant, **2 mol H_2 react with 1 mol O_2 to give 2 mol H_2O**. The coefficients are a ratio of amounts (moles), not of masses: 4.04 g of hydrogen reacts with 32.00 g of oxygen, not 2 g with 1 g.

2.2 What a balanced equation does not tell you

An equation shows the ratio in which substances react. It says nothing about how fast the reaction goes (Reactivity 2.2), whether it goes to completion or reaches equilibrium (Reactivity 2.3), or whether a reactant is actually present in the right amount (the limiting reactant, Booklet 2). For example, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is perfectly balanced, but a real mixture never converts completely to ammonia.

2.3 State symbols

Symbol	Meaning	Use it for
(s)	solid	metals, ionic solids, and **precipitates** that form in solution, e.g. $PbI_2(s)$, $BaSO_4(s)$, $AgCl(s)$
(l)	liquid	pure liquids: water formed in neutralization, $H_2O(l)$; ethanol, $C_2H_5OH(l)$; bromine, $Br_2(l)$
(g)	gas	gases, and water when it forms above 100 °C, e.g. combustion in an engine: $H_2O(g)$
(aq)	aqueous	a substance **dissolved in water**: acids such as $HCl(aq)$, alkalis such as $NaOH(aq)$, and soluble salts

Two words in a question tell you which reactants to write. **"Burning"** or **"combustion"** means that O_2 is a reactant. **"Heating"** or **"thermal decomposition"** does not: heat only supplies the energy, and the compound breaks down on its own, as in $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$.

2.4 Balancing equations

Method — balancing an equation by inspection

1. Write correct formulas for every reactant and product. Remember the diatomic elements: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2.
2. Start with the most complex species. Balance a polyatomic ion (SO_4^{2-}, NO_3^-, CO_3^{2-}) as a single unit if it appears unchanged on both sides.
3. Leave elements that appear on their own (O_2, H_2, Cl_2, a metal) until last, because changing their coefficient affects nothing else.
4. **Only ever change coefficients, never subscripts.** Changing H_2O to H_2O_2 makes a different substance.
5. A fraction such as $\frac{1}{2}O_2$ is allowed while you work. Multiply through to whole numbers if the question asks for them.
6. Check every element, and the charges if there are ions, on both sides.

Worked example 1 — from words to balanced equations with state symbols

Reaction	Skeleton equation	Balanced, with state symbols
calcium carbonate and dilute hydrochloric acid	$CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$	2 Cl needed on the left: **$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$**
aluminium burns in oxygen	$Al + O_2 \rightarrow Al_2O_3$	O: 2 and 3 → lowest common multiple 6: **$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$**
sodium reacts with water	$Na + H_2O \rightarrow NaOH + H_2$	H: 2 on the left, 3 on the right → double H_2O and NaOH: **$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$**
lead(II) nitrate solution and potassium iodide solution: a yellow precipitate	$Pb(NO_3)_2 + KI \rightarrow PbI_2 + KNO_3$	treat NO_3 as a unit; 2 I needed: **$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$**
thermal decomposition of copper(II) nitrate	$Cu(NO_3)_2 \rightarrow CuO + NO_2 + O_2$	N: 2 → $2NO_2$; O: 6 = 1 + 4 + 1 → $\frac{1}{2}O_2$; double: **$2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$**

Worked example 2 — harder equations: fractions and the algebraic method

Q. Balance (a) $NH_3 + O_2 \rightarrow NO + H_2O$ and (b) $KMnO_4 + HCl \rightarrow KCl + MnCl_2 + Cl_2 + H_2O$. ★

(a) N: 1 NH_3 → 1 NO. H: 3 in NH_3 → 1½ H_2O. O on the right: 1 + 1½ = 2½ → 1¼ O_2.

$NH_3 + 1\frac{1}{4}O_2 \rightarrow NO + 1\frac{1}{2}H_2O$; multiply by 4: **$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$**

(b) Let the coefficients be a $KMnO_4$ + b HCl → c KCl + d $MnCl_2$ + e Cl_2 + f H_2O.

K: $a = c$ · Mn: $a = d$ · O: $4a = f$ · H: $b = 2f = 8a$ · Cl: $b = c + 2d + 2e \rightarrow 8a = 3a + 2e \rightarrow e = 2.5a$

Put $a = 2$: **$2KMnO_4 + 16HCl \rightarrow 2KCl + 2MnCl_2 + 5Cl_2 + 8H_2O$**

Check: Cl on the right = 2 + 4 + 10 = 16. The algebraic method always works, even when inspection gets stuck.

2.5 Ionic equations and half-equations

Some ions take no part in a reaction: they are **spectator ions**. Leaving them out gives an **ionic equation** that shows only the change that actually happens. In the precipitation in worked example 1, K^+ and NO_3^- are spectators:

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s) \quad \text{and for any neutralization:} \quad H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

A redox reaction can be split further into two **half-equations**, one for the species that loses electrons (oxidation) and one for the species that gains them (reduction). For magnesium and hydrochloric acid, $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$ splits into $Mg \rightarrow Mg^{2+} + 2e^-$ and $2H^+ + 2e^- \rightarrow H_2$. Half-equations are useful when the two processes happen in different places, as in electrochemical cells and fuel cells, and when you need to work out the mole ratio in a redox titration (worked example 17). Both balance atoms **and** charge. They are developed fully in Reactivity 3.2.

Examiner traps in 2.1.1

- Changing a subscript to balance ($H_2O \rightarrow H_2O_2$, or $Al_2O_3 \rightarrow AlO_3$). Only coefficients may change.
- Writing Cl, O or H as single atoms. As elements they are $Cl_2(g)$, $O_2(g)$ and $H_2(g)$.
- Giving water the symbol (aq). Water is the solvent: it is $H_2O(l)$. Solutions are (aq), precipitates (s).
- Dropping the brackets of a polyatomic ion: $Mg(NO_3)_2$ is not $MgNO_{32}$. The subscript outside the bracket multiplies everything inside.
- Adding O_2 to a "heating" question. Thermal decomposition has no oxygen reactant.
- Omitting state symbols. The syllabus says to include them, and mark schemes often require them.

3 Reactivity 2.1.2 — The mole toolkit SL + HL

The coefficients in an equation give the ratio of *amounts* in moles. Mass, gas volume and concentration are what we actually measure, so every calculation has three steps: convert what you are given into moles, use the mole ratio, and convert the answer back into what is asked for. The tools come from Structure 1.4 and 1.5.

Definitions to learn (from Structure 1.4)

The **amount of substance**, n, is measured in moles. One mole contains 6.02×10^{23} particles (the Avogadro constant, N_A).

The **molar mass**, M, is the mass of one mole of a substance, in g mol^{-1}. It is the sum of the A_r values in the formula.

The **molar concentration**, c, is the amount of solute per unit volume of solution, in mol dm^{-3}.

The **molar volume**, V_m, is the volume occupied by one mole of any gas at a given temperature and pressure: 22.7 dm^3 mol^{-1} at STP.

To find moles from…	Formula	Units	Rearranged
a mass	$n = m / M$	m in g, M in g mol^{-1}	$m = nM$
a gas volume at STP	$n = V / V_m$	V in dm^3, V_m = 22.7 dm^3 mol^{-1}	$V = nV_m$
a gas volume at any T and p	$pV = nRT$	p in Pa, V in m^3, T in K	$n = pV / RT$
a solution	$n = cV$	c in mol dm^{-3}, V in dm^3	$c = n / V$
a number of particles	$n = N / N_A$	$N_A = 6.02 \times 10^{23}$ mol^{-1}	$N = nN_A$

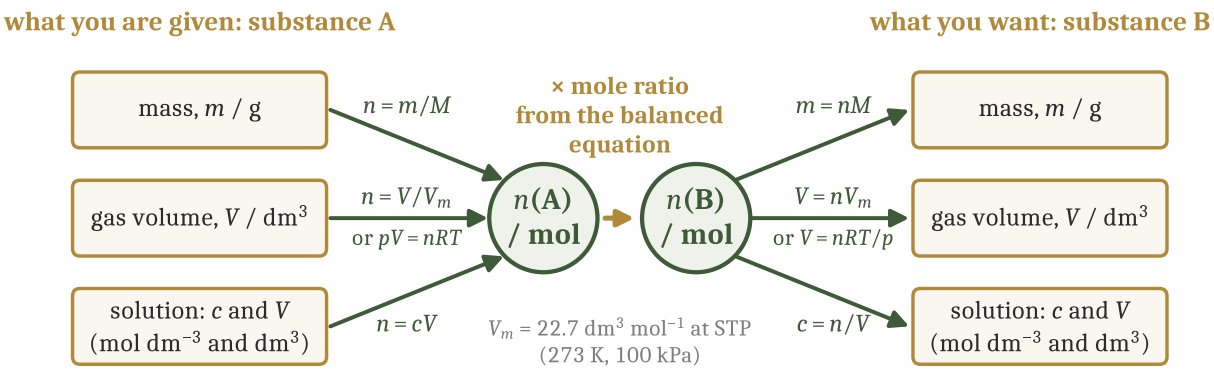


Figure 1 — The mole bridge. Convert whatever you are given into moles of A, cross the bridge with the mole ratio, then convert moles of B into whatever is asked for.

Method — every reacting-quantities calculation

1. Write the balanced equation (with state symbols if asked).
2. Calculate the amount, in mol, of the substance you know everything about.
3. Use the coefficients: n(B) = n(A) × (coefficient of B ÷ coefficient of A).
4. Convert n(B) into the quantity asked for: mass, volume or concentration.

Work only with the two substances in the question. Converting every species in the equation into grams wastes time and invites mistakes.

3.2 Units and significant figures

- **Volume:** cm^3 ÷ 1000 → dm^3; dm^3 ÷ 1000 → m^3. So 25.00 cm^3 = 0.02500 dm^3 = 2.500 × 10^{-5} m^3. A litre (L) is 1 dm^3 and mL is cm^3, but IB questions use cm^3, dm^3 and m^3.
- **Mass:** 1 kg = 10^3 g; 1 tonne = 10^3 kg = 10^6 g. **Temperature:** T/K = t/°C + 273. **Pressure:** 1 kPa = 10^3 Pa.
- **Molar masses:** use the data booklet A_r values to two decimal places, e.g. M(CO_2) = 12.01 + 2(16.00) = 44.01 g mol^{-1}.
- **Significant figures:** give the final answer to the same number of significant figures as the least precise data value. Keep at least one extra figure in intermediate steps and round only at the end.

Worked example 3 — warming up with the toolkit

Question	Working	Answer
amount in 4.50 g of water	$M(H_2O) = 2(1.01) + 16.00 = 18.02$; 4.50 / 18.02	**0.250 mol**
amount of NaCl in 350.0 cm^3 of 0.150 mol dm^{-3} solution	0.150 × (350.0 / 1000) = 0.150 × 0.3500	**0.0525 mol**
amount in 1.20 dm^3 of CO_2 at STP	1.20 / 22.7	**0.0529 mol**
volume of 0.0400 mol O_2 at STP	0.0400 × 22.7 = 0.908 dm^3	**908 cm^3**
0.250 mol dm^{-3} NaOH in g dm^{-3}	M(NaOH) = 22.99 + 16.00 + 1.01 = 40.00; 0.250 × 40.00	**10.0 g dm^{-3}**

4 Reactivity 2.1.2 — Reacting masses SL + HL

4.1 Masses from the mole ratio

When methane burns, $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$, the mole ratio 1 : 2 : 1 : 2 becomes a mass ratio once each amount is multiplied by its molar mass: 16.05 g + 64.00 g → 44.01 g + 36.04 g. Both sides total 80.05 g. That is conservation of mass again, and a useful check on any equation.

Worked example 4 — the mass of a product: two equivalent methods

Q. Calculate the masses of carbon dioxide and water formed when 5.00 g of propane, C_3H_8, burns completely.

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$ · $M(C_3H_8) = 3(12.01) + 8(1.01) = 44.11$ g mol^{-1}

Mole method: $n(C_3H_8) = 5.00 / 44.11 = 0.1134$ mol

$n(CO_2) = 3 \times 0.1134 = 0.3401$ mol; $m = 0.3401 \times 44.01 =$ **15.0 g CO_2**

$n(H_2O) = 4 \times 0.1134 = 0.4534$ mol; $m = 0.4534 \times 18.02 =$ **8.17 g H_2O**

Ratio method: 1 mol C_3H_8 (44.11 g) gives 3 mol CO_2 (132.03 g), so 5.00 g gives 5.00 × 132.03 / 44.11 = 15.0 g

Both methods give the same answer. Use whichever you find more reliable, but always show the mole ratio.

Worked example 5 — the mass of a reactant needed

Q. In the thermite reaction, $Fe_2O_3(s) + 2Al(s) \rightarrow Al_2O_3(s) + 2Fe(l)$. Calculate the mass of aluminium needed to react with 10.0 g of iron(III) oxide, and the mass of iron produced.

$M(Fe_2O_3) = 2(55.85) + 3(16.00) = 159.70$ g mol^{-1}; $n(Fe_2O_3) = 10.0 / 159.70 = 0.06262$ mol

$n(Al) = 2 \times 0.06262 = 0.1252$ mol; $m(Al) = 0.1252 \times 26.98 =$ **3.38 g**

$n(Fe) = 2 \times 0.06262 = 0.1252$ mol; $m(Fe) = 0.1252 \times 55.85 =$ **6.99 g**

Worked example 6 — industrial quantities in tonnes

Q. Calculate the mass of hydrogen needed to make 1.00 tonne of ammonia: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$.

Via moles: 1.00 t = 1.00×10^6 g; $n(NH_3) = 1.00 \times 10^6 / 17.04 = 5.869 \times 10^4$ mol

$n(H_2) = (3/2) \times 5.869 \times 10^4 = 8.803 \times 10^4$ mol; $m = 8.803 \times 10^4 \times 2.02 = 1.78 \times 10^5$ g = **0.178 t (178 kg)**

Faster, by mass ratio: 3 mol H_2 (6.06 g) make 2 mol NH_3 (34.08 g), so 1.00 t × 6.06 / 34.08 = 0.178 t

A mass ratio is the same in any unit, so tonnes can stay as tonnes. Moles cannot: a "tonne-mole" is not a unit you can use with g mol^{-1}.

4.2 Percentage purity and composition

Real samples are rarely pure. If only one component reacts to give a product you can weigh or measure, the mole ratio gives the mass of that component, and

percentage purity = (mass of the pure substance ÷ mass of the sample) × 100

Worked example 7 — purity of limestone

Q. 2.50 g of impure limestone was heated strongly to constant mass. It lost 0.968 g of carbon dioxide. Calculate the percentage by mass of calcium carbonate in the limestone.

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ · $n(CO_2) = 0.968 / 44.01 = 0.02200$ mol

$n(CaCO_3) = n(CO_2) = 0.02200$ mol; $m = 0.02200 \times 100.09 = 2.201$ g

percentage of $CaCO_3$ = 2.201 / 2.50 × 100 = **88.1 %**

Model answer: "State two assumptions made in calculating the percentage of calcium carbonate in the limestone." (2 marks)

All of the calcium carbonate decomposed (it was heated to constant mass) [1]. The impurities did not decompose or lose mass on heating, so all of the mass lost was carbon dioxide [1].

Worked example 8 — water of crystallisation

Q. 2.46 g of hydrated magnesium sulfate, $MgSO_4 \cdot xH_2O$, was heated to constant mass. 1.20 g of anhydrous $MgSO_4$ remained. Determine x.

$MgSO_4 \cdot xH_2O(s) \rightarrow MgSO_4(s) + xH_2O(g)$ · mass of water lost = 2.46 − 1.20 = 1.26 g

$n(MgSO_4) = 1.20 / 120.38 = 9.97 \times 10^{-3}$ mol; $n(H_2O) = 1.26 / 18.02 = 6.99 \times 10^{-2}$ mol

$x = 6.99 \times 10^{-2} / 9.97 \times 10^{-3} = 7.01 \rightarrow$ **$x = 7$, $MgSO_4 \cdot 7H_2O$**

Heating to constant mass makes sure that all of the water has been driven off. If it has not, the mass of water is too small and x comes out too low.

Worked example 9 — a multi-step process ★

Q. Sulfuric acid is made in three steps: $S(s) + O_2(g) \rightarrow SO_2(g)$; $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$; $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$. Assuming complete conversion at each step, calculate the mass of sulfuric acid made from 1.00 tonne of sulfur, and the mass of 98.0 % acid this represents.

Follow the sulfur: $1\ S \rightarrow 1\ SO_2 \rightarrow 1\ SO_3 \rightarrow 1\ H_2SO_4$, so the overall ratio is 1 : 1

$M(S) = 32.07$; $M(H_2SO_4) = 2(1.01) + 32.07 + 4(16.00) = 98.09\ g\ mol^{-1}$

mass of $H_2SO_4 = 1.00\ t \times 98.09 / 32.07 =$ **3.06 t**; mass of 98.0 % acid = 3.06 / 0.980 = **3.12 t**

For a chain of reactions, combine the ratios into one overall ratio, keeping track of the element that is carried through.

Examiner traps with reacting masses

- Using the coefficients as a *mass* ratio ("2 g of Al for every 1 g of Fe_2O_3"). Coefficients are a mole ratio.
- Multiplying the molar mass by the coefficient *and* then using the mole ratio as well, so the coefficient is counted twice.
- Rounding A_r values (C = 12, O = 16) when the question expects the data booklet values to two decimal places.
- Rounding intermediate answers to two figures, then losing the third significant figure in the final answer.
- Getting a percentage purity above 100 %: usually a sign that the wrong mole ratio or molar mass was used.

5 Reactivity 2.1.2 — Reacting volumes of gases SL + HL

5.1 Avogadro's law

Definition to learn

Avogadro's law: equal volumes of all gases, measured at the same temperature and pressure, contain equal numbers of molecules (so $V \propto n$).

It follows that for gases measured under the same conditions, **the ratio of volumes is the same as the ratio of moles**. You can read reacting volumes straight from the coefficients without calculating any moles at all. This only applies to gases: solids and liquids take up almost no volume by comparison.

Avogadro proposed his hypothesis in 1811, but it was not accepted for nearly 50 years, partly because it clashed with Dalton's belief that gaseous elements were single atoms. Cannizzaro showed in 1860 how it could be used to find molecular masses. The law *describes* how gases behave, but it does not *explain* it. The explanation comes from the kinetic model of an ideal gas (Structure 1.5): the particles are so far apart that their own size, and the forces between them, are negligible. So the volume depends only on the number of particles, the temperature and the pressure, and not on what the gas is.

5.2 Volume ratios

Worked example 10 — reacting volumes

Q. 40 cm^3 of propane is burned in 300 cm^3 of oxygen (an excess). All volumes are measured at the same pressure. Calculate the volume of oxygen used and of carbon dioxide formed, and the total volume of gas (a) at 120 °C and (b) after cooling to 25 °C.

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O$ volume ratio 1 : 5 : 3 : 4

O_2 used = 5 × 40 = **200 cm^3** (100 cm^3 remains); CO_2 formed = 3 × 40 = **120 cm^3**

(a) at 120 °C water is a gas: 120 + 160 + 100 = **380 cm^3** (b) at 25 °C water is a liquid: 120 + 100 = **220 cm^3**

Volumes are compared at the same temperature and pressure, so a gas volume "at 120 °C" is compared with the other gases at 120 °C. Figure 2 shows the result.

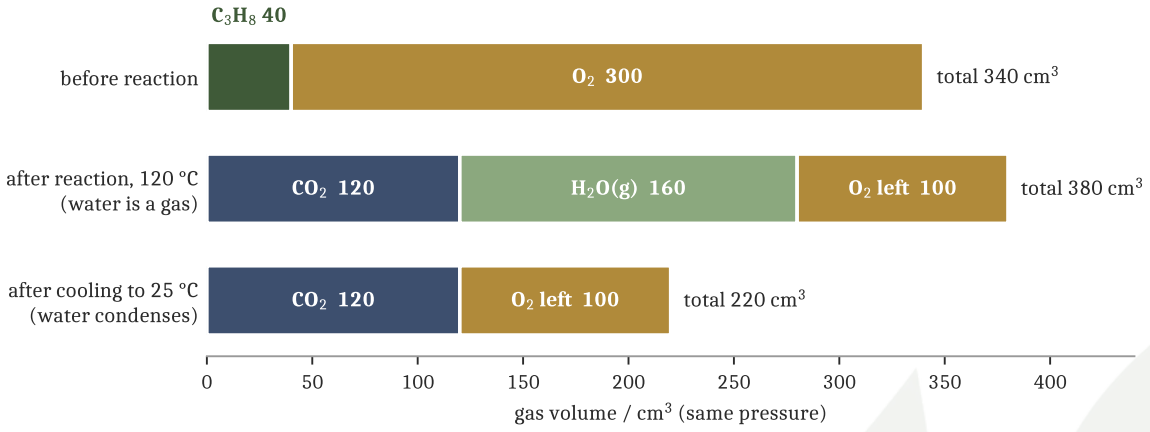


Figure 2 — Worked example 10. Because volume ratio = mole ratio, the coefficients can be applied directly to gas volumes. Cooling removes the water vapour; passing the gas through NaOH(aq) would then remove the acidic CO_2, leaving only the excess O_2.

Worked example 11 — the formula of a gaseous hydrocarbon from volume data ★

Q. 20 cm^3 of a gaseous hydrocarbon, C_xH_y, was mixed with 150 cm^3 of oxygen (an excess) and exploded. After cooling to room temperature the gas volume was 110 cm^3. After passing it through sodium hydroxide solution, which absorbs CO_2, 50 cm^3 remained. Determine the formula of the hydrocarbon.

$C_xH_y + (x + y/4)O_2 \rightarrow xCO_2 + (y/2)H_2O(l)$

volume of CO_2 = 110 − 50 = 60 cm^3, so 20 cm^3 gives 60 cm^3: x = 60 / 20 = 3

O_2 remaining = 50 cm^3, so O_2 used = 150 − 50 = 100 cm^3: $x + y/4$ = 100 / 20 = 5

3 + $y/4$ = 5 → y = 8: **C_3H_8, propane**

This is how chemists once found molecular formulas. It is an application of Avogadro's law, so it only works because the hydrocarbon, oxygen and carbon dioxide are all gases.

5.3 The molar volume at STP

By Avogadro's law, one mole of any gas occupies the same volume under the same conditions. At **STP**, defined as 273 K (0 °C) and 100 kPa, the molar volume is **22.7 $dm^3\ mol^{-1}$** (2.27 × 10^{-2} $m^3\ mol^{-1}$). Both values are in the data booklet. STP is not the same as the thermodynamic standard state of Reactivity 1, which specifies 100 kPa but not 273 K.

Worked example 12 — a mass to a gas volume

Q. Calculate the volume of hydrogen, at STP, produced when 0.540 g of aluminium reacts with excess hydrochloric acid.

$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$

$n(Al) = 0.540 / 26.98 = 0.02001$ mol; $n(H_2) = (3/2) \times 0.02001 = 0.03002$ mol

$V(H_2) = 0.03002 \times 22.7 = 0.682$ dm^3 = **682 cm^3**

5.4 Other temperatures and pressures

Molar volume is not a constant. Heating a gas at constant pressure makes the particles move faster and spread out, so the molar volume **increases**. Raising the pressure at constant temperature pushes them closer together, so it **decreases**. At 298 K and 100 kPa, for example, V_m is about 24.8 dm^3 mol^{-1}. Under any conditions other than STP, use the ideal gas equation from Structure 1.5, $pV = nRT$, in SI units: p in Pa, V in m^3, T in K.

Worked example 13 — a gas volume away from STP

Q. 2.45 g of potassium chlorate(V) is heated until it is fully decomposed: $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$. Calculate the volume of oxygen produced at 25 °C and 101 kPa.

$M(KClO_3) = 39.10 + 35.45 + 3(16.00) = 122.55$ g mol^{-1}; $n(KClO_3) = 2.45 / 122.55 = 0.01999$ mol

$n(O_2) = (3/2) \times 0.01999 = 0.02999$ mol · $T = 25 + 273 = 298$ K; $p = 1.01 \times 10^5$ Pa

$V = nRT / p = 0.02999 \times 8.31 \times 298 / 1.01 \times 10^5 = 7.35 \times 10^{-4}$ m^3 = **735 cm^3**

At STP the same amount would occupy $0.02999 \times 22.7 = 0.681$ dm^3. The warmer gas takes up more room.

Model answer: "State Avogadro's law and use it to explain why 10 cm^3 of nitrogen reacts with hydrogen to form 20 cm^3 of ammonia, all volumes measured at the same temperature and pressure." (2 marks)

Equal volumes of all gases at the same temperature and pressure contain equal numbers of molecules [1]. So the volume ratio equals the mole ratio in $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$: 1 volume of N_2 gives 2 volumes of NH_3, so 10 cm^3 gives 20 cm^3 [1].

Examiner traps with gas volumes

- Using 22.7 dm^3 mol^{-1} when the conditions are not STP. Check the temperature and pressure first.
- Using cm^3 or kPa in $pV = nRT$. With $R = 8.31$, the units must be Pa, m^3 and K. A common slip is that 1 dm^3 = 10^{-3} m^3, not 10^{-6}.
- Applying volume ratios to a liquid or solid, e.g. counting liquid water at room temperature as a gas volume.
- Comparing volumes measured under different conditions. Volume ratio = mole ratio only at the same T and p.
- Confusing STP (273 K, 100 kPa) with standard state conditions in thermochemistry (100 kPa, usually 298 K).

6 Reactivity 2.1.2 — Solutions and titration SL + HL

6.1 Concentration and standard solutions

For a reaction in solution, the amount of each solute comes from $n = cV$, with the volume in dm^3. Concentration can also be given as a **mass concentration** in g dm^{-3}, which is $c \times M$. Diluting a solution adds solvent but not solute, so the amount of solute is unchanged: $c_1V_1 = c_2V_2$.

Definitions to learn

A **standard solution** is a solution of accurately known concentration. It is made by dissolving an accurately weighed solid and making the solution up to the mark in a volumetric flask.

Titration is a technique of volumetric analysis in which one solution is added from a burette to a measured volume (an **aliquot**) of another until the reaction is just complete.

The **equivalence point** is when the reactants have reacted exactly in the mole ratio of the equation. The **end point** is when the indicator changes colour. A good indicator has its end point at the equivalence point.

The **titre** is the volume of solution added from the burette (final reading − initial reading). **Concordant** titres agree closely with each other, usually to within 0.10 cm^3 (some schools use 0.05 cm^3).

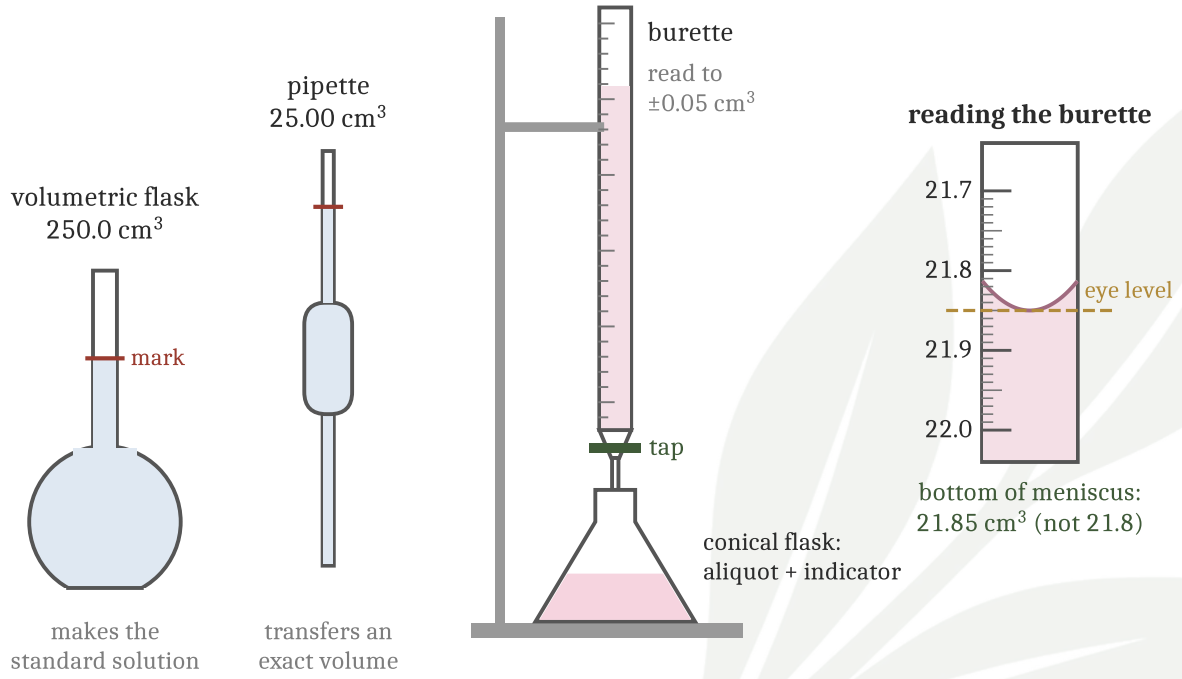


Figure 3 — Titration apparatus. The standard solution is made in a volumetric flask; a pipette transfers an exact aliquot into the conical flask; the burette delivers the other solution until the indicator changes colour. Read the burette at eye level, at the bottom of the meniscus, to the nearest 0.05 cm^3.

6.2 The volumetric analysis method

Method — titration calculations

1. Write the balanced equation, if it is not given. Note the mole ratio: a diprotic acid such as H_2SO_4 reacts with 2 mol of NaOH.
2. Find the amount of the reactant whose concentration and volume you both know: $n = cV$ (volume in dm^3).
3. Use the mole ratio to find the amount of the other reactant.
4. Divide by its volume in dm^3: $c = n / V$. If only an aliquot was titrated, scale the amount up to the whole solution first.

6.3 Recording and choosing titres

A 50 cm^3 burette has 0.1 cm^3 divisions. The uncertainty of an analogue scale is half the smallest division, so each reading is recorded to **two decimal places**, ending in 0 or 5 (0.00, 12.35, 23.50; never 0, 12.3 or 23.5). A titre is the difference of two readings, so its uncertainty is ±0.05 + 0.05 = **±0.10 cm^3**. The first, rough titre shows roughly where the end point is and is not used. Accurate titrations are repeated until you have concordant titres, and the mean of the concordant titres is used in the calculation.

Worked example 14 — titration results to a concentration

Q. 25.00 cm^3 portions of sodium hydroxide solution were titrated with 0.0500 mol dm^{-3} sulfuric acid. Use the results to calculate the concentration of the sodium hydroxide.

Titration	rough	1	2	3
final burette reading / cm^3 ± 0.05	22.40	22.05	22.90	22.20
initial burette reading / cm^3 ± 0.05	0.00	0.20	1.10	0.35
titre / cm^3 ± 0.10	22.40	21.85	21.80	21.85

concordant titres: 21.85, 21.80, 21.85; mean = 21.83 cm^3 (the rough titre is ignored)

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

$n(H_2SO_4) = 0.0500 \times 0.02183 = 1.092 \times 10^{-3}$ mol; $n(NaOH) = 2 \times 1.092 \times 10^{-3} = 2.183 \times 10^{-3}$ mol

$c(NaOH) = 2.183 \times 10^{-3} / 0.02500 =$ **0.0873 mol dm^{-3}**

Worked example 15 — a standard solution and an aliquot

Q. 1.325 g of anhydrous sodium carbonate was dissolved in water and made up to 250.0 cm^3 in a volumetric flask. A 25.00 cm^3 aliquot needed 23.45 cm^3 of hydrochloric acid for neutralization. Calculate the concentration of the acid.

$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$ · $M(Na_2CO_3) = 105.99$ g mol^{-1}

$n(Na_2CO_3)$ in the flask = 1.325 / 105.99 = 1.250×10^{-2} mol, so $c = 1.250 \times 10^{-2} / 0.2500 = 0.05000$ mol dm^{-3}

$n(Na_2CO_3)$ in 25.00 cm^3 = $0.05000 \times 0.02500 = 1.250 \times 10^{-3}$ mol (one tenth of the flask)

$n(HCl) = 2 \times 1.250 \times 10^{-3} = 2.500 \times 10^{-3}$ mol; $c(HCl) = 2.500 \times 10^{-3} / 0.02345 =$ **0.107 mol dm^{-3}**

Worked example 16 — solutions that make a precipitate or a gas

Q. (a) Calculate the mass of barium sulfate precipitated when excess sodium sulfate solution is added to 50.0 cm^3 of 0.200 mol dm^{-3} barium chloride, and the minimum volume of 0.150 mol dm^{-3} sodium sulfate needed. (b) Calculate the mass of magnesium that reacts with 25.0 cm^3 of 2.00 mol dm^{-3} hydrochloric acid, and the volume of hydrogen formed at STP.

(a) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$ · $n(BaCl_2) = 0.200 \times 0.0500 = 0.0100$ mol

$n(BaSO_4) = 0.0100$ mol; $m = 0.0100 \times 233.40 =$ **2.33 g**; $V(Na_2SO_4) = 0.0100 / 0.150 = 0.0667\ dm^3 =$ **66.7 cm^3**

(b) $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$ · $n(HCl) = 2.00 \times 0.0250 = 0.0500$ mol

$n(Mg) = 0.0500 / 2 = 0.0250$ mol; $m = 0.0250 \times 24.31 =$ **0.608 g**; $V(H_2) = 0.0250 \times 22.7 =$ **0.568 dm^3**

Worked example 17 — a redox titration (link to Reactivity 3.2) ★

Q. An iron tablet of mass 1.250 g was dissolved in dilute sulfuric acid and made up to 100.0 cm^3. 10.00 cm^3 portions were titrated with 0.0100 mol dm^{-3} potassium manganate(VII); the mean titre was 14.20 cm^3. The ionic equation is $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$. Calculate the percentage by mass of iron in the tablet.

$n(MnO_4^-) = 0.0100 \times 0.01420 = 1.420 \times 10^{-4}$ mol

$n(Fe^{2+})$ in 10.00 $cm^3 = 5 \times 1.420 \times 10^{-4} = 7.100 \times 10^{-4}$ mol

in 100.0 cm^3 (× 10): 7.100×10^{-3} mol; $m(Fe) = 7.100 \times 10^{-3} \times 55.85 = 0.3965$ g

percentage of iron = 0.3965 / 1.250 × 100 = **31.7 %**

The mole ratio 1 : 5 comes from electrons: Mn goes from +7 to +2, gaining 5 e^-, while each Fe^{2+} loses only 1 e^-. MnO_4^- is purple and Mn^{2+} almost colourless, so the manganate(VII) acts as its own indicator: the end point is the first permanent pale pink.

Examiner traps in titration calculations

- Using cm^3 in $n = cV$. Divide by 1000 first: 25.00 cm^3 = 0.02500 dm^3.
- Forgetting the mole ratio for a diprotic acid: H_2SO_4 needs **2** NaOH; Na_2CO_3 needs **2** HCl.
- Forgetting the scale factor when an aliquot (25.00 cm^3 of 250.0 cm^3) was titrated: the flask holds 10 times as much.
- Including the rough titre, or a titre that is not concordant, in the mean.
- Recording burette readings as 22.4 or 0 instead of 22.40 and 0.00.
- Giving the titre uncertainty as ±0.05 cm^3. Two readings make ±0.10 cm^3.

7 Reactivity 2.1.2 — Back titration SL + HL

7.1 When a direct titration will not work

In a **back titration** a known **excess** of one reagent is added to the sample. When the reaction is complete, the reagent that is left over is titrated with a standard solution. Subtracting the excess from the total gives the amount that reacted with the sample. It is used when:

- the sample is an **insoluble solid** ($CaCO_3$ in an antacid tablet, eggshell or limestone) that cannot be put in a burette or pipette;
- the reaction with the sample is **slow**, so the end point of a direct titration would be hard to judge;
- the substance being analysed is **volatile**, such as ammonia driven off from an ammonium salt by warm alkali;
- the sample contains **impurities** or is coloured, so a direct end point is hard to see.

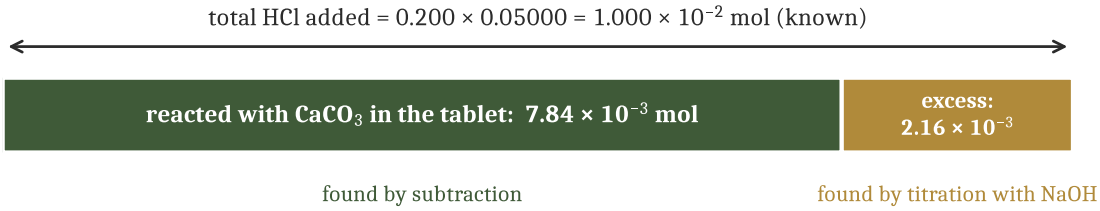


Figure 4 — The logic of a back titration, using the numbers in worked example 18. The total is known; the excess is measured; the amount that reacted is found by subtraction.

Method — back titration calculations

1. Total amount of reagent added: $n(\text{total}) = cV$.
2. Amount left over (excess) from the titration, using its own equation and mole ratio. Scale up if only an aliquot was titrated.
3. Amount that reacted with the sample = $n(\text{total}) - n(\text{excess})$.
4. Use the equation for the reaction with the sample to find the amount, then the mass and percentage, of the substance analysed.

Worked example 18 — calcium carbonate in an antacid tablet

Q. A 0.500 g antacid tablet containing calcium carbonate was added to 50.00 cm^3 of 0.200 mol dm^{-3} hydrochloric acid (an excess). When the fizzing had stopped, the remaining acid needed 21.60 cm^3 of 0.100 mol dm^{-3} sodium hydroxide for neutralization. Calculate the percentage by mass of calcium carbonate in the tablet.

total HCl added = $0.200 \times 0.05000 = 1.000 \times 10^{-2}$ mol

excess HCl: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$, so $n(HCl) = n(NaOH) = 0.100 \times 0.02160 = 2.160 \times 10^{-3}$ mol

HCl that reacted with the tablet = $1.000 \times 10^{-2} - 2.160 \times 10^{-3} = 7.840 \times 10^{-3}$ mol

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$: $n(CaCO_3) = 7.840 \times 10^{-3} / 2 = 3.920 \times 10^{-3}$ mol

$m(CaCO_3) = 3.920 \times 10^{-3} \times 100.09 = 0.3924$ g; percentage = 0.3924 / 0.500 × 100 = **78.5 %**

The calculation assumes that only the calcium carbonate reacts with the acid (no other basic ingredient), and that all of it reacted.

Model answer: "Explain why a back titration, rather than a direct titration, is used to find the percentage of calcium carbonate in an antacid tablet." (2 marks)

Calcium carbonate is an insoluble solid, so it cannot be made into a solution of known volume for a direct titration [1]. It also reacts slowly with acid, so the end point would be hard to judge; adding a known excess of acid lets it react completely, and the excess can then be titrated accurately [1].

Examiner traps in back titrations

- Using the excess as the amount that reacted with the sample. The amount that reacted is total − excess.
- Using the wrong equation at each stage: one mole ratio for the titration, another for the reaction with the sample.
- Forgetting to scale the excess up when the reaction mixture was made up to 250.0 cm^3 and only 25.00 cm^3 was titrated.
- Quoting a result without its assumptions, for example that no other component of the tablet reacts with acid.

8 Key points to memorise

Everything below is either wording that examiners look for or knowledge you must carry into the exam yourself.

A Definitions

- **Balanced equation:** the same number of atoms of each element, and the same total charge, on both sides; the coefficients give the mole ratio.
- **State symbols:** (s), (l), (g), (aq). Water is (l), a solution is (aq), a precipitate is (s).
- **Avogadro's law:** equal volumes of all gases at the same temperature and pressure contain equal numbers of molecules.
- **Molar volume:** the volume of one mole of gas; 22.7 $dm^3\ mol^{-1}$ at STP (273 K, 100 kPa).
- **Standard solution:** a solution of accurately known concentration.
- **Equivalence point:** reactants in exact mole ratio. **End point:** indicator changes colour. **Titre:** volume added from the burette.
- **Back titration:** a known excess of reagent is added and the excess is titrated; amount reacted = total − excess.

B Equations

- $n = m / M$ · $n = V / V_m$ (STP) · $pV = nRT$ (Pa, m^3, K) · $n = cV$ (dm^3) · $N = nN_A$
- Mole ratio: n(B) = n(A) × (coefficient of B / coefficient of A). For gases at the same T and p, volume ratio = mole ratio.
- Dilution: $c_1V_1 = c_2V_2$. Mass concentration (g dm^{-3}) = $c \times M$.
- Percentage purity = mass of pure substance / mass of sample × 100.
- Titre uncertainty = ±0.10 cm^3 (two readings of ±0.05 cm^3). Record burette readings to two decimal places.

C Explanations you will be asked for

- Equations must balance because mass is conserved: atoms are rearranged, not created or destroyed. Change coefficients, never subscripts.
- Gas volumes react in the ratio of the coefficients because equal volumes contain equal numbers of molecules (gases only, same T and p).
- Molar volume increases with temperature and decreases with pressure; use $pV = nRT$ away from STP.
- Titrations are repeated to obtain concordant titres; the rough titre is not used.
- A back titration is used for insoluble, slow-reacting or volatile analytes, or when the end point is hard to see.
- Purity calculations assume that the impurities do not react and that the reaction (or decomposition) is complete.

9 Check yourself

Use A_r values from the data booklet (H 1.01, Li 6.94, C 12.01, N 14.01, O 16.00, Na 22.99, Mg 24.31, Al 26.98, S 32.07, Cl 35.45, K 39.10, Ca 40.08, Ti 47.87, Mn 54.94, Fe 55.85, Cu 63.55, Zn 65.38, Ag 107.87, Ba 137.33), V_m = 22.7 dm^3 mol^{-1} at STP and R = 8.31 J K^{-1} mol^{-1}. Marks are shown in brackets; allow about 1.2 minutes per mark. Full answers are in section 10.

Part A — Multiple choice SL + HL

1. What is the sum of the coefficients when the equation for the complete combustion of butane is balanced using the smallest whole numbers? $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$

A 23 **B** 29 **C** 33 **D** 36

2. Which equation is correctly balanced?

A $Fe + Cl_2 \rightarrow FeCl_3$
B $2Fe + 3Cl_2 \rightarrow 2FeCl_3$
C $2Fe + 3Cl \rightarrow 2FeCl_3$
D $Fe + 3Cl_2 \rightarrow FeCl_6$

3. Ammonia burns in oxygen to form nitrogen and water. What is the coefficient of O_2 in the balanced equation with the smallest whole numbers?

A 2 **B** 3 **C** 5 **D** 6

4. Which equation, with state symbols, represents the reaction of solid magnesium carbonate with dilute nitric acid?

A $MgCO_3(aq) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(aq) + CO_2(g)$
B $MgCO_3(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
C $MgCO_3(s) + 2HNO_3(l) \rightarrow Mg(NO_3)_2(s) + H_2O(l) + CO_2(g)$
D $MgCO_3(s) + HNO_3(aq) \rightarrow MgNO_3(aq) + H_2O(l) + CO_2(g)$

5. What mass of carbon dioxide forms when 0.200 mol of ethanol, C_2H_5OH, burns completely?

A 8.80 g **B** 17.6 g **C** 26.4 g **D** 35.2 g

6. What mass of magnesium oxide forms when 1.215 g of magnesium burns completely? $2Mg + O_2 \rightarrow 2MgO$

A 0.806 g **B** 1.01 g **C** 2.01 g **D** 4.03 g

7. Which sample contains the greatest number of molecules?

A 2.02 g of H_2
B 11.35 dm^3 of O_2 at STP
C 44.01 g of CO_2
D 34.08 g of NH_3

8. What volume of oxygen is needed to burn 20 cm^3 of ethene, C_2H_4, completely? All volumes are at the same temperature and pressure.

A 20 cm^3 **B** 40 cm^3 **C** 60 cm^3 **D** 80 cm^3

9. What is the volume of 0.0500 mol of any ideal gas at STP?

A 1.14 dm^3 **B** 11.4 dm^3 **C** 454 dm^3 **D** 1.14 cm^3

10. How does the molar volume of a gas change when its temperature is increased at constant pressure?

A It increases.
B It decreases.
C It stays the same.
D It depends on the identity of the gas.

11. What amount of sodium hydroxide is in 25.0 cm^3 of a 0.400 mol dm^{-3} solution?

A 1.00×10^{-3} mol **B** 0.0100 mol **C** 0.100 mol **D** 10.0 mol

12. What volume of 0.100 mol dm^{-3} H_2SO_4 neutralizes 20.0 cm^3 of 0.150 mol dm^{-3} NaOH?

A 7.50 cm^3 **B** 15.0 cm^3 **C** 30.0 cm^3 **D** 60.0 cm^3

13. A student records titres of 24.60 (rough), 24.15, 24.40, 24.20 and 24.15 cm^3. Which titres should be used to calculate the mean?

A all five
B 24.15, 24.20 and 24.15
C 24.40 and 24.20
D 24.60 and 24.40

14. Each burette reading has an uncertainty of ±0.05 cm^3. What is the uncertainty of a titre?

A ±0.025 cm^3 **B** ±0.05 cm^3 **C** ±0.10 cm^3 **D** ±0.20 cm^3

15. Why is a back titration used to analyse the calcium carbonate in eggshell?

A Calcium carbonate is an insoluble solid that reacts slowly with acid.
B Calcium carbonate is a strong acid.
C No indicator changes colour with hydrochloric acid.
D A back titration does not need a standard solution.

16. 10.0 cm^3 of a 1.00 mol dm^{-3} solution is diluted with water to 250.0 cm^3. What is the new concentration?

A 0.0250 mol dm^{-3} **B** 0.0400 mol dm^{-3} **C** 0.400 mol dm^{-3} **D** 25.0 mol dm^{-3}

Part B — Structured questions SL + HL

17. Write balanced equations, with state symbols, for: [8]

(a) zinc and dilute hydrochloric acid; (b) calcium hydroxide solution and nitric acid; (c) the reduction of iron(III) oxide by carbon monoxide; (d) aluminium and copper(II) sulfate solution, forming aluminium sulfate and copper; (e) sodium carbonate solution and hydrochloric acid; (f) the thermal decomposition of sodium nitrate to sodium nitrite, $NaNO_2$, and oxygen; (g) silver nitrate solution and calcium chloride solution; (h) white phosphorus, P_4, burning in chlorine to form phosphorus(V) chloride.

18. ★ Balance, using the smallest whole numbers: (a) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$; (b) $FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$; (c) $K_2Cr_2O_7 + HCl \rightarrow KCl + CrCl_3 + Cl_2 + H_2O$. [3]

19. Titanium is extracted by reducing titanium(IV) chloride with magnesium: $TiCl_4(g) + 2Mg(l) \rightarrow Ti(s) + 2MgCl_2(l)$.

(a) Calculate the mass of magnesium needed to react with 1.00 kg of $TiCl_4$. [2]

(b) Calculate the mass of titanium produced from 1.00 kg of $TiCl_4$. [1]

(c) Calculate the mass of $TiCl_4$, in tonnes, needed to produce 1.00 tonne of titanium. [2]

20. 5.00 g of hydrated sodium carbonate, $Na_2CO_3 \cdot xH_2O$, was heated to constant mass, leaving 1.85 g of anhydrous Na_2CO_3.

(a) Determine the value of x. [3]

(b) Explain why the solid is heated to constant mass. [1]

21. A 3.00 g sample of impure sodium hydrogencarbonate lost 0.954 g when heated to constant mass: $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$.

(a) Calculate the percentage by mass of $NaHCO_3$ in the sample. [3]

(b) State one assumption you have made. [1]

22. ★ Nitric acid is made from ammonia in three steps: $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$; $2NO + O_2 \rightarrow 2NO_2$; $4NO_2 + O_2 + 2H_2O \rightarrow 4HNO_3$. Assuming complete conversion, calculate the mass of HNO_3 made from 1.00 tonne of ammonia. [3]

23. 25 cm^3 of butane, C_4H_{10}, is burned in 250 cm^3 of oxygen. All volumes are measured at the same temperature and pressure.

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

(b) Calculate the volume of oxygen used and the volume of carbon dioxide formed. [2]

(c) Calculate the total volume of gas after cooling to room temperature. [1]

(d) Calculate the volume of gas left after the mixture in (c) is passed through NaOH(aq). [1]

24. ★ 10 cm^3 of a gaseous hydrocarbon was exploded with 80 cm^3 of oxygen (an excess). After cooling to room temperature the volume was 65 cm^3; after passing through NaOH(aq) it was 25 cm^3. Determine the molecular formula of the hydrocarbon and suggest one possible structure. [4]

25. Answer at STP.

(a) Calculate the volume occupied by 3.20 g of oxygen gas. [1]

(b) Deduce which contains more molecules: 5.00 dm^3 of nitrogen or 6.50 g of oxygen. [2]

(c) Calculate the mass of 500 cm^3 of carbon dioxide. [2]

26. In spacecraft, lithium hydroxide removes exhaled carbon dioxide: $2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O(l)$.

(a) Calculate the volume of CO_2, measured at STP, absorbed by 1.00 kg of LiOH. [2]

(b) Calculate the volume of this CO_2 in the cabin, at 22 °C and 101 kPa. [2]

27. ★ 25.0 cm^3 of hydrogen peroxide solution decomposed completely, $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$, giving 84.5 cm^3 of oxygen at 20 °C and 102 kPa. Calculate the concentration of the hydrogen peroxide solution in mol dm^{-3}. [4]

28. 25.00 cm^3 of vinegar was diluted to 250.0 cm^3 in a volumetric flask. 25.00 cm^3 portions of the diluted vinegar were titrated with 0.100 mol dm^{-3} NaOH.

Titration	rough	1	2	3	4
final reading / cm^3 ± 0.05	21.90	21.95	43.55	21.60	22.50
initial reading / cm^3 ± 0.05	0.00	0.50	21.95	0.20	1.05

(a) Calculate each titre, identify the concordant titres and calculate their mean. [1]

(b) Write the equation for the reaction of ethanoic acid, CH_3COOH, with NaOH. [1]

(c) Calculate the amount of NaOH in the mean titre. [1]

(d) Calculate the concentration of ethanoic acid in the diluted vinegar. [1]

(e) Calculate the concentration of ethanoic acid in the original vinegar, in mol dm^{-3} and in g dm^{-3}. [2]

29. A standard solution was made by dissolving 1.575 g of ethanedioic acid crystals, $H_2C_2O_4 \cdot 2H_2O$, and making the solution up to 250.0 cm^3. 25.00 cm^3 portions needed a mean titre of 24.80 cm^3 of NaOH(aq): $H_2C_2O_4(aq) + 2NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O(l)$.

(a) Calculate the concentration of the ethanedioic acid solution. [2]

(b) Calculate the concentration of the sodium hydroxide solution. [2]

(c) Suggest why solid sodium hydroxide is not weighed out to make a standard solution directly. [1]

30. 25.0 cm^3 of 0.150 mol dm^{-3} silver nitrate solution is added to excess sodium chloride solution.

(a) Write the ionic equation, with state symbols, for the precipitation. [1]

(b) Calculate the mass of precipitate formed. [2]

(c) Calculate the minimum volume of 0.100 mol dm^{-3} $CaCl_2(aq)$ that would precipitate all of the silver ions. [1]

31. 0.350 g of zinc reacts completely with excess dilute sulfuric acid.

(a) Write the equation, with state symbols. [1]

(b) Calculate the volume of hydrogen produced at STP. [2]

(c) Calculate the minimum volume of 0.500 mol dm^{-3} sulfuric acid needed. [1]

32. ★ 25.00 cm^3 of hydrogen peroxide solution was acidified and titrated with 0.0200 mol dm^{-3} $KMnO_4$; the mean titre was 22.40 cm^3. $2MnO_4^-(aq) + 5H_2O_2(aq) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5O_2(g) + 8H_2O(l)$

(a) Calculate the amount of MnO_4^- in the mean titre. [1]

(b) Calculate the amount of H_2O_2 in 25.00 cm^3. [1]

(c) Calculate the concentration of the H_2O_2 solution. [1]

(d) Calculate the volume of oxygen, at STP, that 1.00 dm^3 of this solution releases when it decomposes: $2H_2O_2 \rightarrow 2H_2O + O_2$. [2]

33. ★ A 1.20 g sample of a fertilizer containing ammonium sulfate was boiled with 50.00 cm^3 of 0.500 mol dm^{-3} NaOH (an excess) until no more ammonia was given off: $(NH_4)_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2NH_3 + 2H_2O$. The remaining NaOH needed 28.40 cm^3 of 0.250 mol dm^{-3} HCl for neutralization.

(a) Calculate the percentage by mass of ammonium sulfate in the fertilizer. [5]

(b) Calculate the percentage by mass of nitrogen in the fertilizer. [1]

34. ★ A 0.600 g antacid tablet containing magnesium hydroxide was dissolved in 50.00 cm^3 of 0.500 mol dm^{-3} HCl (an excess). The solution was made up to 250.0 cm^3, and 25.00 cm^3 portions needed 12.60 cm^3 of 0.100 mol dm^{-3} NaOH. $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$. Calculate the percentage by mass of $Mg(OH)_2$ in the tablet. [5]

35. ★ A 25.00 cm^3 solution containing both Na_2CO_3 and $NaHCO_3$ was titrated with 0.100 mol dm^{-3} HCl. The phenolphthalein end point was reached after 11.20 cm^3 ($Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$). Methyl orange was then added and a further 27.50 cm^3 was needed ($NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$).

Calculate the concentrations of Na_2CO_3 and $NaHCO_3$ in the original solution. [4]

Part A: 16 marks · Part B: 88 marks · Total: 104 marks (about 2 hours)

10 Answers

Part A

1. C. $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$: 2 + 13 + 8 + 10 = 33.

2. B. A is unbalanced; C writes chlorine as atoms; D has the wrong formula (changing subscripts).

3. B. $4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$.

4. B. Water is (l), the acid and the soluble salt are (aq); magnesium nitrate is $Mg(NO_3)_2$, so $2HNO_3$ are needed.

5. B. $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$: 0.400 mol CO_2 × 44.01 = 17.6 g.

6. C. n(Mg) = 1.215/24.31 = 0.04998 mol = n(MgO); × 40.31 = 2.01 g.

7. D. A and C are 1.00 mol, B is 0.500 mol, D is 2.00 mol.

8. C. $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$, so 3 × 20 = 60 cm^3.

9. A. 0.0500 × 22.7 = 1.135 dm^3.

10. A. The particles move faster and occupy more space at the same pressure; for an ideal gas the identity does not matter.

11. B. 0.400 × 0.0250 = 0.0100 mol.

12. B. n(NaOH) = 3.00 × 10^{-3} mol → $n(H_2SO_4)$ = 1.50 × 10^{-3} mol → 1.50 × 10^{-3}/0.100 = 0.0150 dm^3.

13. B. The rough titre is not used, and 24.40 is not concordant. Mean = 24.17 cm^3.

14. C. A titre is the difference of two readings: ±0.05 + 0.05.

15. A.

16. B. c_2 = 1.00 × 10.0/250.0 = 0.0400 mol dm^{-3}.

Part B

17. One mark each (formulas, balancing and state symbols): (a) $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ · (b) $Ca(OH)_2(aq) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$ · (c) $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$ · (d) $2Al(s) + 3CuSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$ · (e) $Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$ · (f) $2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$ · (g) $2AgNO_3(aq) + CaCl_2(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$ · (h) $P_4(s) + 10Cl_2(g) \rightarrow 4PCl_5(s)$.

18. (a) $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$ [1] · (b) $4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$ [1] · (c) $K_2Cr_2O_7 + 14HCl \rightarrow 2KCl + 2CrCl_3 + 3Cl_2 + 7H_2O$ [1].

19. (a) M($TiCl_4$) = 189.67; n = 1000/189.67 = 5.272 mol [1]; n(Mg) = 10.54 mol × 24.31 = 256 g [1]. (b) 5.272 × 47.87 = 252 g [1]. (c) 189.67 g of $TiCl_4$ gives 47.87 g of Ti [1], so 1.00 × 189.67/47.87 = 3.96 t [1].

20. (a) mass of water = 5.00 − 1.85 = 3.15 g [1]; $n(Na_2CO_3)$ = 1.85/105.99 = 0.01745 mol; $n(H_2O)$ = 3.15/18.02 = 0.1748 mol [1]; ratio = 10.0, so x = 10 [1]. (b) To make sure that all of the water of crystallisation has been removed [1].

21. (a) 2 mol $NaHCO_3$ (168.02 g) lose 1 mol H_2O + 1 mol CO_2 = 62.03 g [1]; mass of $NaHCO_3$ = 0.954 × 168.02/62.03 = 2.58 g [1]; percentage = 2.58/3.00 × 100 = 86.1 % [1]. (b) The impurities do not decompose or lose mass on heating (or: the decomposition is complete) [1].

22. Each step is 1 : 1 in nitrogen, so 1 mol NH_3 → 1 mol HNO_3 [1]; M(NH_3) = 17.04, M(HNO_3) = 63.02 [1]; mass = 1.00 × 63.02/17.04 = 3.70 t [1].

23. (a) $C_4H_{10}(g) + 6½O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$ (or $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$) [1]. (b) O_2 used = 6.5 × 25 = 162.5 cm^3 [1]; CO_2 = 4 × 25 = 100 cm^3 [1]. (c) water condenses: 100 + (250 − 162.5) = 187.5 cm^3 [1]. (d) CO_2 is absorbed: 87.5 cm^3 of oxygen remains [1].

24. CO_2 = 65 − 25 = 40 cm^3, so x = 4 [1]; O_2 used = 80 − 25 = 55 cm^3 [1]; $x + y/4$ = 5.5, so y = 6: C_4H_6 [1]; e.g. buta-1,3-diene, $CH_2{=}CH{-}CH{=}CH_2$ (or but-1-yne, $HC{\equiv}C{-}CH_2{-}CH_3$, or but-2-yne) [1].

25. (a) 3.20/32.00 = 0.100 mol; × 22.7 = 2.27 dm^3 [1]. (b) $n(N_2)$ = 5.00/22.7 = 0.220 mol; $n(O_2)$ = 6.50/32.00 = 0.203 mol [1]; the nitrogen contains more molecules [1]. (c) n = 0.500/22.7 = 0.0220 mol [1]; m = 0.0220 × 44.01 = 0.969 g [1].

26. (a) $n(\text{LiOH}) = 1000/23.95 = 41.75$ mol, so $n(CO_2) = 20.88$ mol [1]; $V = 20.88 \times 22.7 = 474$ dm^3 [1]. (b) $V = nRT/p = 20.88 \times 8.31 \times 295/1.01 \times 10^5$ [1] $= 0.507$ m$^3 = 507$ dm^3 [1].

27. $n(O_2) = pV/RT = (1.02 \times 10^5 \times 8.45 \times 10^{-5})/(8.31 \times 293)$ [1] $= 3.54 \times 10^{-3}$ mol [1]; $n(H_2O_2) = 2 \times 3.54 \times 10^{-3} = 7.08 \times 10^{-3}$ mol [1]; $c = 7.08 \times 10^{-3}/0.0250 = 0.283$ mol dm^{-3} [1].

28. (a) titres 21.90 (rough), 21.45, 21.60, 21.40, 21.45; concordant: 21.45, 21.40, 21.45; mean = 21.43 cm^3 [1]. (b) $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$ [1]. (c) $0.100 \times 0.02143 = 2.143 \times 10^{-3}$ mol [1]. (d) $2.143 \times 10^{-3}/0.02500 = 0.0857$ mol dm^{-3} [1]. (e) $\times 10 = 0.857$ mol dm^{-3} [1]; $\times 60.06 = 51.5$ g dm^{-3} [1].

29. (a) $M(H_2C_2O_4\cdot 2H_2O) = 126.08$; $n = 1.575/126.08 = 1.249 \times 10^{-2}$ mol [1]; $c = 1.249 \times 10^{-2}/0.2500 = 0.0500$ mol dm^{-3} [1]. (b) n(acid) in 25.00 cm$^3 = 1.249 \times 10^{-3}$ mol, so $n(\text{NaOH}) = 2.498 \times 10^{-3}$ mol [1]; $c = 2.498 \times 10^{-3}/0.02480 = 0.101$ mol dm^{-3} [1]. (c) Solid NaOH absorbs water vapour and carbon dioxide from the air, so its weighed mass is not an accurate measure of the amount of NaOH [1].

30. (a) $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ [1]. (b) $n = 0.150 \times 0.0250 = 3.75 \times 10^{-3}$ mol [1]; $\times 143.32 = 0.537$ g [1]. (c) each $CaCl_2$ gives $2Cl^-$: $n(CaCl_2) = 1.875 \times 10^{-3}$ mol; $V = 1.875 \times 10^{-3}/0.100 = 0.0188$ dm$^3 = 18.8$ cm^3 [1].

31. (a) $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$ [1]. (b) $n(\text{Zn}) = 0.350/65.38 = 5.35 \times 10^{-3}$ mol $= n(H_2)$ [1]; $V = 5.35 \times 10^{-3} \times 22.7 = 0.122$ dm^3 (122 cm^3) [1]. (c) $5.35 \times 10^{-3}/0.500 = 1.07 \times 10^{-2}$ dm$^3 = 10.7$ cm^3 [1].

32. (a) $0.0200 \times 0.02240 = 4.48 \times 10^{-4}$ mol [1]. (b) $\times 5/2 = 1.12 \times 10^{-3}$ mol [1]. (c) $1.12 \times 10^{-3}/0.02500 = 0.0448$ mol dm^{-3} [1]. (d) 1.00 dm^3 contains 0.0448 mol H_2O_2, giving 0.0224 mol O_2 [1]; $V = 0.0224 \times 22.7 = 0.508$ dm^3 [1].

33. (a) total NaOH $= 0.500 \times 0.05000 = 0.0250$ mol [1]; excess NaOH $= n(\text{HCl}) = 0.250 \times 0.02840 = 7.10 \times 10^{-3}$ mol [1]; NaOH that reacted $= 0.0250 - 0.00710 = 0.0179$ mol [1]; $n((NH_4)_2SO_4) = 0.0179/2 = 8.95 \times 10^{-3}$ mol [1]; $m = 8.95 \times 10^{-3} \times 132.17 = 1.18$ g, so 98.6 % [1]. (b) one NH_3 (one N) per NaOH reacted: $n(\text{N}) = 0.0179$ mol; $0.0179 \times 14.01/1.20 \times 100 = 20.9$ % [1].

34. total HCl $= 0.500 \times 0.05000 = 0.0250$ mol [1]; excess in 25.00 cm$^3 = 0.100 \times 0.01260 = 1.26 \times 10^{-3}$ mol, so 1.26×10^{-2} mol in 250.0 cm^3 [1]; HCl reacted $= 0.0250 - 0.0126 = 0.0124$ mol [1]; $n(Mg(OH)_2) = 6.20 \times 10^{-3}$ mol [1]; $m = 6.20 \times 10^{-3} \times 58.33 = 0.362$ g, so 60.3 % [1].

35. $n(Na_2CO_3) = n$(HCl, step 1) $= 0.100 \times 0.01120 = 1.12 \times 10^{-3}$ mol, so $c(Na_2CO_3) = 0.0448$ mol dm^{-3} [1]; HCl in step 2 $= 2.75 \times 10^{-3}$ mol reacts with all the $NaHCO_3$, including the 1.12×10^{-3} mol made in step 1 [1]; original $n(NaHCO_3) = 2.75 \times 10^{-3} - 1.12 \times 10^{-3} = 1.63 \times 10^{-3}$ mol [1]; $c(NaHCO_3) = 0.0652$ mol dm^{-3} [1].

Where this fits in Reactivity 2

- **Before this:** Structure 1.4 (the mole, molar mass, concentration) and Structure 1.5 (ideal gases, $pV = nRT$) supply the toolkit.
- **Booklet 1 (this booklet)** — Reactivity 2.1.1–2.1.2: equations, reacting masses, gas volumes, solutions and titrations.
- **Booklet 2** — Reactivity 2.1.3–2.1.5: limiting and excess reactants, theoretical and percentage yield, and atom economy.
- **Next:** Reactivity 2.2 — how fast? Rates of reaction.

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