MEGA LECTURE

IB DP CHEMISTRY

Reactivity 2: How Much, How Fast and How Far?

Reactivity 2.1 How Much? The Amount of Chemical Change

Revision Notes · Standard and Higher Level

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Original notes prepared for the IB Diploma Programme Chemistry course (2025 syllabus)

What the syllabus requires

Reactivity 2.1 is the quantitative heart of the course: it asks *how much* product a reaction can make and how efficiently. Everything here rests on the mole concept (Structure 1.4). Use this checklist before the exam.

Understanding	You should be able to...
Balanced equations	Write and balance full, ionic and net ionic equations, with correct state symbols and spectator ions.
Mole ratios	Read the stoichiometric ratio straight from the balanced equation and use it to convert between amounts.
Limiting reactant	Identify the limiting and excess reactants, and calculate the amount of product and of excess left over.
Yield	Calculate theoretical yield, and percentage yield from an actual yield.
Atom economy	Calculate percentage atom economy and link it to green chemistry.
Volumetric analysis	Carry out titration calculations, including back titrations, to find an unknown concentration or molar mass.
Mass and gas	Relate the mass or the volume of a gaseous product to the amount of reactant.

> **Exam note:** The whole of R2.1 is common to SL and HL. Back titrations and multi-step conversions appear more often at HL and in Paper 2 data questions. Sections flagged **(HL)** show the more demanding treatment.

1. Chemical equations

A balanced chemical equation is a quantitative statement. The formulae tell you *what* reacts; the coefficients tell you *in what ratio of amounts*. Because atoms are conserved (Structure 1.4), every element must appear in equal numbers on both sides, and the total charge must balance too.

Balancing: a systematic method

- Write correct formulae first and never change a subscript to balance → only coefficients may be adjusted.
- Balance metals and polyatomic ions first, then non-metals, then hydrogen, then oxygen last.
- Treat spectator polyatomic ions (SO_4^{2-}, NO_3^-) as single units.
- Finish by clearing fractions to give the smallest whole-number ratio.

State symbols

Add (s) solid, (l) liquid, (g) gas and (aq) aqueous (dissolved in water). They are not decoration: they decide whether an ionic equation can be written and identify which species is the precipitate or the gas.

Worked example 1 — balancing a combustion equation

Balance the complete combustion of propane, C_3H_8.

Unbalanced: $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

Balance C (3 → 3 CO_2), then H (8 → 4 H_2O). Oxygen on the right = (3×2) + (4×1) = 10, so 5 O_2.

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$ Check: C 3=3, H 8=8, O 10=10.

Ionic and net ionic equations

When ionic compounds dissolve they dissociate into free ions. An **ionic equation** shows every aqueous species as separate ions. Ions that appear unchanged on both sides are **spectator ions**; cancelling them leaves the **net ionic equation**, which shows the chemical change that actually happens.

Worked example 2 — ionic and net ionic equations

Silver nitrate solution is mixed with sodium chloride solution; a white precipitate forms.

Full: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Ionic: $Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$

Na^+ and NO_3^- are spectators. Net ionic:

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ Charge balances: (+1) + (−1) = 0 on the left, 0 on the right.

2. Mole ratios from balanced equations

The coefficients in a balanced equation are the **mole ratio** (stoichiometric ratio) of the species. This single idea powers every calculation in this topic. The universal route for any quantitative problem is:

known quantity → moles of known → (× mole ratio) moles of wanted → wanted quantity

Recall the conversions that get you into and out of moles: $n = m / M$ for mass, $n = c \times V$ for solutions, and $n = V / V_m$ for gases (molar volume V_m = 22.7 $dm^3\ mol^{-1}$ at STP in the 2025 data booklet).

Worked example 3 — using a mole ratio

In the Haber process, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. How many moles of hydrogen react with 0.50 mol of nitrogen, and how much ammonia forms?

Ratio $N_2 : H_2 : NH_3 = 1 : 3 : 2$.

$n(H_2) = 0.50 \times 3 =$ **1.5 mol**. $n(NH_3) = 0.50 \times 2 =$ **1.0 mol**.

> **Watch the ratio:** The mole ratio is a ratio of *amounts*, never of masses. Always convert masses (or volumes) to moles before applying it, then convert back at the end.

Reacting masses

The commonest version of this route converts a mass of reactant into a mass of product. The three steps are always the same: mass → moles (divide by M), apply the mole ratio, then moles → mass (multiply by M). Only the two molar masses and the ratio change from question to question.

Worked example 3b — mass of reactant to mass of product

What mass of aluminium oxide forms when 5.40 g of aluminium burns completely in oxygen? $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$. ($A_r$ Al 27.0; M Al_2O_3 102.0)

n(Al) = 5.40 / 27.0 = 0.200 mol.

Ratio Al : Al_2O_3 = 4 : 2 = 2 : 1, so $n(Al_2O_3)$ = 0.200 / 2 = 0.100 mol.

mass = 0.100 × 102.0 = **10.2 g**.

3. Limiting and excess reactants

Reactants are rarely mixed in exactly the stoichiometric ratio. The **limiting reactant** is the one that runs out first; it determines the maximum amount of product. The others are in **excess** and some is left over when the reaction stops.

How to find the limiting reactant

- Convert the amount of each reactant to moles.
- Divide each by its coefficient in the balanced equation.
- The smallest value is the limiting reactant. Base all product amounts on it.
- Excess remaining = initial moles − moles that reacted.

Worked example 4 — limiting reactant and excess

13.08 g of zinc (A_r 65.38) is added to 300 cm^3 of 1.0 mol dm^{-3} hydrochloric acid. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$.

n(Zn) = 13.08 / 65.38 = 0.20 mol. n(HCl) = 1.0 × (300/1000) = 0.30 mol.

Divide by coefficients: Zn → 0.20/1 = 0.20; HCl → 0.30/2 = 0.15. HCl is smaller, so **HCl is limiting**.

$n(H_2)$ = 0.30 × (1/2) = **0.15 mol** (volume at STP = 0.15 × 22.7 = 3.4 dm^3).

Zn reacting = 0.30 × (1/2) = 0.15 mol, so Zn left over = 0.20 − 0.15 = **0.05 mol** = 3.3 g.

Worked example 5 — limiting reactant from two amounts

2.0 mol of N_2 is mixed with 3.0 mol of H_2: $N_2 + 3H_2 \rightarrow 2NH_3$. Find the amount of NH_3 and the excess remaining.

Divide by coefficients: $N_2 \rightarrow 2.0/1 = 2.0$; $H_2 \rightarrow 3.0/3 = 1.0$. **$H_2$ is limiting**.

$n(NH_3) = 3.0 \times (2/3) =$ **2.0 mol**.

N_2 reacting = $3.0 \times (1/3) = 1.0$ mol, so N_2 left = $2.0 - 1.0 =$ **1.0 mol**.

Worked example 5b — limiting reactant giving mass of product

8.10 g of zinc oxide (M 81.4) is warmed with 14.7 g of sulfuric acid (M 98.1): $ZnO + H_2SO_4 \rightarrow ZnSO_4 + H_2O$. Find the limiting reactant and the mass of $ZnSO_4$ formed. (M $ZnSO_4$ 161.5)

$n(ZnO) = 8.10 / 81.4 = 0.0995$ mol. $n(H_2SO_4) = 14.7 / 98.1 = 0.150$ mol.

Ratio is 1 : 1, so ZnO (0.0995 mol) is the smaller amount and is **limiting**; acid is in excess.

$n(ZnSO_4) = 0.0995$ mol; mass = $0.0995 \times 161.5 =$ **16.1 g**.

H_2SO_4 left over = $0.150 - 0.0995 = 0.0505$ mol.

Common error: Do not decide the limiting reactant just from which mass is smaller. You must divide moles by the coefficient. A reactant present in the larger amount can still be limiting if its coefficient is large.

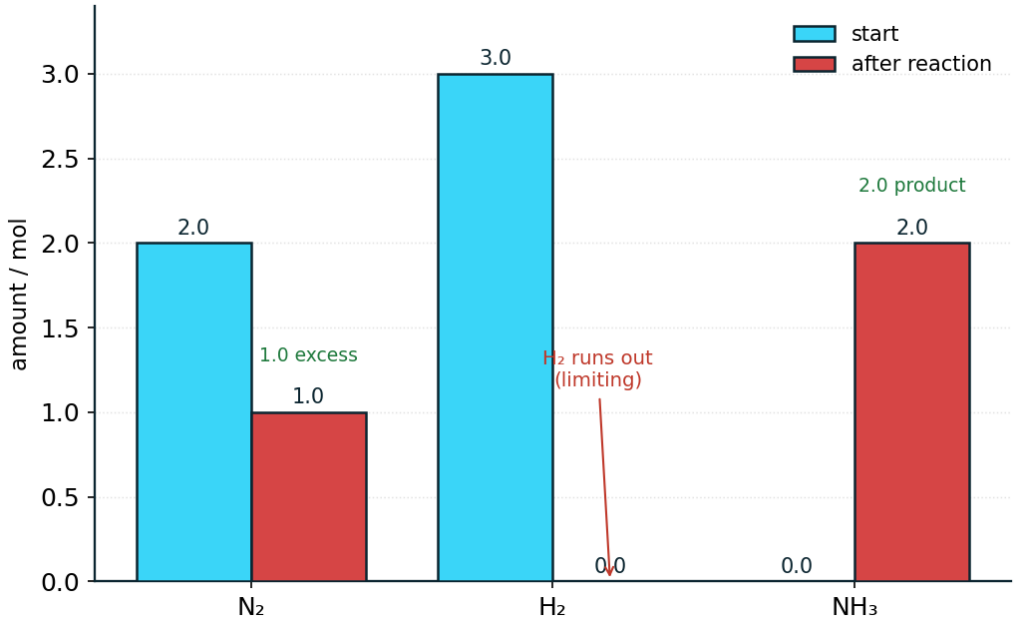

Figure. Reaction inventory for $N_2 + 3H_2 \rightarrow 2NH_3$ starting from 2.0 mol N_2 and 3.0 mol H_2. H_2 runs out first (limiting), giving 2.0 mol NH_3 and leaving 1.0 mol N_2 in excess.

4. Yield and atom economy

Two different efficiency measures are examined. **Percentage yield** asks how much of the possible product you actually obtained; **atom economy** asks how much of the reactant mass ends up in the useful product.

Theoretical, actual and percentage yield

The **theoretical yield** is the maximum mass calculable from the limiting reactant, assuming the reaction goes to completion. The **actual yield** is what is measured. Real yields are lower because reactions may be incomplete or reversible, side reactions occur, and product is lost on filtering, transfer and purification.

$$\textbf{percentage yield} = (\textbf{actual yield} / \textbf{theoretical yield}) \times 100\%$$

Worked example 6 — percentage yield

25.0 g of calcium carbonate is heated: $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$. 12.0 g of calcium oxide is obtained. Find the percentage yield. (M: $CaCO_3$ 100.1, CaO 56.1)

$n(CaCO_3) = 25.0 / 100.1 = 0.2498$ mol, so n(CaO) theoretical = 0.2498 mol.

Theoretical mass CaO = 0.2498 × 56.1 = 14.01 g.

Percentage yield = (12.0 / 14.01) × 100 = **85.6%**.

Atom economy

Atom economy measures the proportion of reactant atoms that become useful product. A high atom economy means less waste, which is a central goal of green chemistry.

$$\textbf{\% atom economy} = (\textbf{M}_r \textbf{ of desired product / total M}_r \textbf{ of all products}) \times 100\%$$

Addition and combination reactions (A + B → C) have 100% atom economy because there is only one product. Reactions that also make a by-product have a lower atom economy even if the percentage yield is high → the two measures are independent.

Worked example 7 — atom economy

For $CaCO_3 \rightarrow CaO + CO_2$, find the atom economy for making CaO. (M: CaO 56.1, CO_2 44.0)

Total product mass (per mole) = 56.1 + 44.0 = 100.1.

Atom economy = (56.1 / 100.1) × 100 = **56.0%** → the CO_2 is discarded, so almost half the mass is 'wasted' atoms.

Worked example 7b — working backwards from percentage yield

A reaction that can theoretically give 6.40 g of product proceeds with a 75.0% yield. What actual mass is obtained, and what mass would be lost?

actual = (% yield / 100) × theoretical = 0.750 × 6.40 = **4.80 g**.

mass 'lost' to incomplete reaction and handling = 6.40 − 4.80 = 1.60 g.

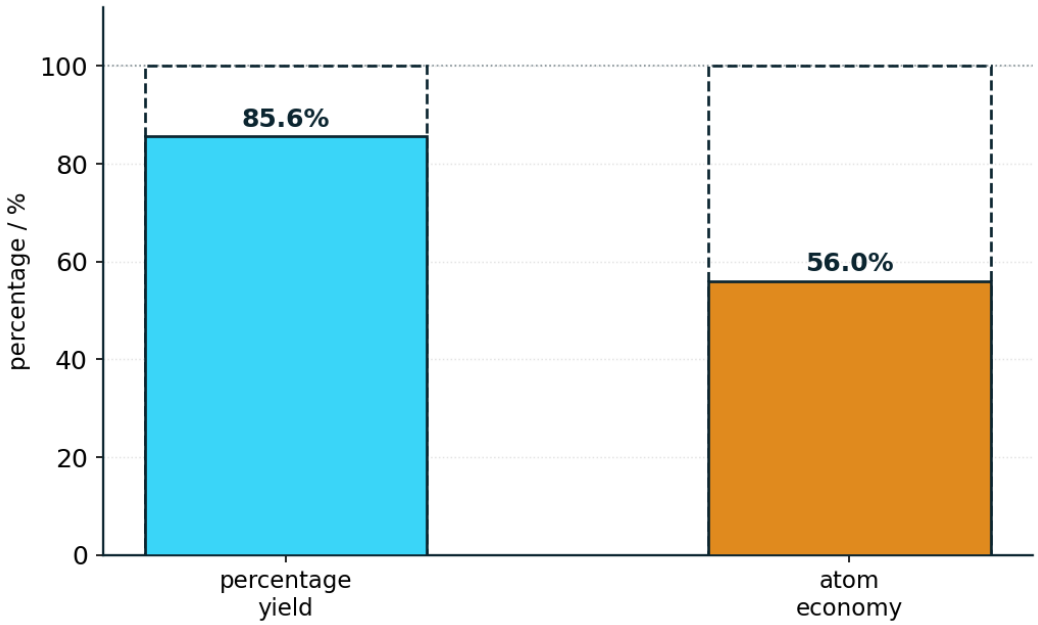

Figure. Two independent efficiency measures for $CaCO_3 \rightarrow CaO + CO_2$: percentage yield = actual/theoretical × 100 = 85.6%, and atom economy = M_r(CaO) / total M_r(products) × 100 = 56.0%.

Comparing the two measures

It is worth being clear how percentage yield and atom economy differ, because exam questions often ask you to distinguish them.

Feature	Percentage yield	Atom economy
What it compares	actual product vs maximum possible	useful product mass vs total reactant mass
Depends on	reaction conditions, losses, side reactions	the balanced equation only
Improved by	optimising conditions, reducing losses	choosing a route with fewer or no by-products
Can it be 100%?	only for a complete, loss-free reaction	only when there is a single product

> **Green chemistry link:** Designing a synthesis with high atom economy reduces raw material use and waste treatment. It is a property of the balanced equation itself, fixed before any practical work, whereas percentage yield depends on the conditions.

5. Volumetric analysis (titrations)

A **titration** measures the volume of one solution that exactly reacts with a known volume of another, so that an unknown concentration or molar mass can be found. In an acid–base titration an indicator (or a pH meter) marks the equivalence point where the acid and base have reacted in their stoichiometric ratio.

Standard solutions

A titration needs one solution of accurately known concentration, called a **standard solution**. It is prepared by dissolving a weighed mass of a pure solid (a primary standard) and making the solution up to a precise volume in a volumetric flask. Its concentration follows directly from $c = n / V$. Concentration is usually

quoted in mol dm^{-3} (amount per unit volume); it can also be given in g dm^{-3}, related by c(g dm^{-3}) = c(mol dm^{-3}) × M.

Dilution

A concentrated stock solution is often diluted to a working concentration. Because the amount of solute is unchanged on adding water, $c_1V_1 = c_2V_2$. Rearrange for whichever quantity is unknown, keeping the volume units consistent on both sides.

Worked example 8b — preparing a dilute solution

What volume of 2.00 mol dm^{-3} HCl is needed to make 250 cm^3 of 0.150 mol dm^{-3} HCl?

$c_1V_1 = c_2V_2$: $V_1 = c_2V_2 / c_1$ = (0.150 × 250) / 2.00 = **18.75 cm^3**.

Measure 18.8 cm^3 of stock and make up to 250 cm^3 in a volumetric flask.

The titration calculation route

- From the standard (known) solution: $n = c \times V$, with V in dm^3.
- Use the mole ratio from the balanced equation to get moles of the unknown.
- Divide by the volume (or mass) of the unknown to get its concentration (or M_r).

Unit trap: Convert cm^3 to dm^3 by dividing by 1000 (25.0 cm^3 = 0.0250 dm^3). Forgetting this is the single most common titration mistake and shifts every answer by a factor of 1000.

Worked example 8 — finding an unknown concentration

25.0 cm^3 of sodium hydroxide is neutralised by 22.40 cm^3 of 0.100 mol dm^{-3} hydrochloric acid. $NaOH + HCl \rightarrow NaCl + H_2O$. Find [NaOH].

n(HCl) = 0.100 × (22.40/1000) = 2.240×10^{-3} mol.

Ratio 1:1, so n(NaOH) = 2.240×10^{-3} mol.

[NaOH] = (2.240×10^{-3}) / (25.0/1000) = **0.0896 mol dm^{-3}**.

Worked example 9 — finding a molar mass

1.50 g of a solid diprotic acid H_2X is dissolved and made up to 250 cm^3. A 25.0 cm^3 portion needs 24.0 cm^3 of 0.100 mol dm^{-3} NaOH. $H_2X + 2NaOH \rightarrow Na_2X + 2H_2O$. Find M_r of H_2X.

n(NaOH) = 0.100 × (24.0/1000) = 2.40×10^{-3} mol.

n(H_2X in 25.0 cm^3) = 2.40×10^{-3} / 2 = 1.20×10^{-3} mol.

In the whole 250 cm^3: $1.20 \times 10^{-3} \times 10$ = 0.0120 mol.

M_r = mass / moles = 1.50 / 0.0120 = **125 g mol^{-1}**.

Back titrations

A **back titration** is used when the sample is insoluble, impure, slow to react, or when a direct end point is hard to see. A **known, excess** amount of reagent is added to react fully with the sample; the **unreacted excess** is then titrated. Subtracting the excess from the original amount gives the amount that reacted with the sample.

amount reacted with sample = (initial moles added) − (moles of excess found)

A worked back titration is given in Section 7. Direct titrations can also measure the purity of a soluble sample, as in the next example.

Worked example 9b — purity from a direct titration

A 2.65 g sample of impure sodium carbonate is dissolved and made up to 250 cm^3. A 25.0 cm^3 portion needs 20.0 cm^3 of 0.200 mol dm^{-3} HCl. $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$. Find the percentage purity. (M Na_2CO_3 106.0)

n(HCl) = 0.200 × (20.0/1000) = 4.00×10^{-3} mol.

Ratio Na_2CO_3 : HCl = 1 : 2, so n(Na_2CO_3 in 25.0 cm^3) = 2.00×10^{-3} mol.

In 250 cm^3: $2.00 \times 10^{-3} \times 10 = 0.0200$ mol; mass = 0.0200 × 106.0 = 2.12 g.

% purity = (2.12 / 2.65) × 100 = **80.0%**.

6. Reacting masses and gas volumes

The same mole-ratio route links the amount of a reactant to the mass or the gas volume of a product. In **gravimetric analysis** a product is isolated (often as a precipitate) and weighed; in **gas-collection** experiments the volume of gas evolved is measured, usually over water or in a gas syringe.

Worked example 10 — volume of gas produced

0.243 g of magnesium (A_r 24.3) reacts with excess hydrochloric acid. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$. Find the volume of H_2 at STP.

n(Mg) = 0.243 / 24.3 = 0.0100 mol.

Ratio Mg : H_2 = 1 : 1, so n(H_2) = 0.0100 mol.

V(H_2) = $n \times V_m$ = 0.0100 × 22.7 = 0.227 dm^3 = **227 cm^3** at STP.

Worked example 11 — mass of precipitate (gravimetric)

Excess silver nitrate is added to a solution containing 0.0250 mol of sodium chloride. $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$. Find the mass of precipitate. (M: AgCl 143.3)

Ratio NaCl : AgCl = 1 : 1, so n(AgCl) = 0.0250 mol.

mass = n × M = 0.0250 × 143.3 = **3.58 g**.

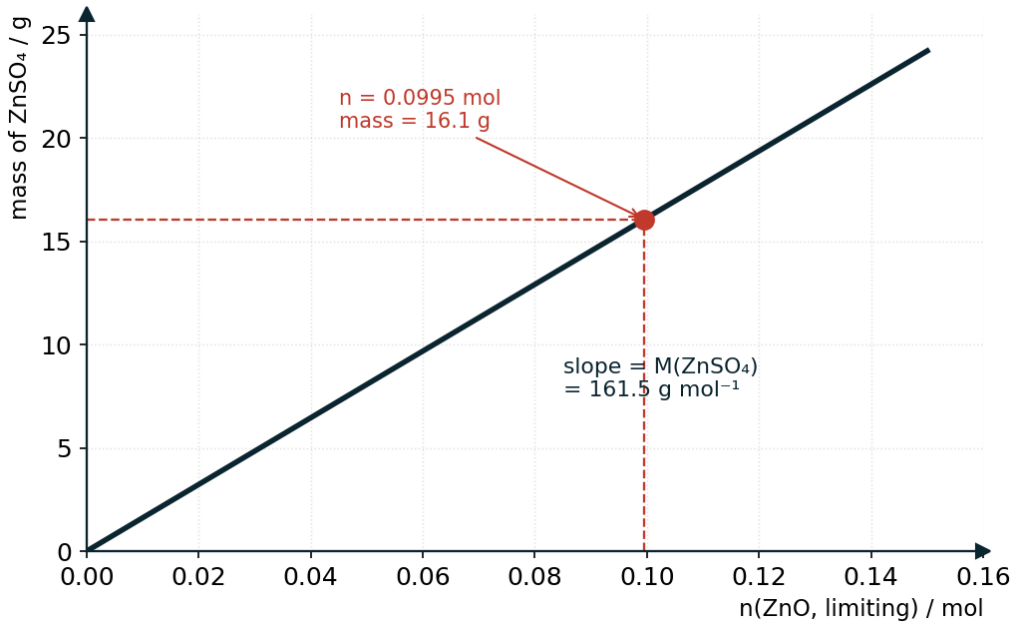

Figure. Mass of $ZnSO_4$ formed is proportional to moles of the limiting reactant ZnO: a straight line through the origin of slope $M(ZnSO_4) = 161.5\ g\ mol^{-1}$. At 0.0995 mol the mass is 16.1 g.

Combining limiting reactant with gas volume

Many data-based questions layer two ideas: first decide the limiting reactant, then use only its amount to find the gas volume or the mass of product. Never base the product on the reactant in excess.

Worked example 11b — limiting reactant then gas volume

0.120 g of magnesium (A_r 24.3) is added to 20.0 cm^3 of 0.300 mol dm^{-3} HCl. $Mg + 2HCl \rightarrow MgCl_2 + H_2$. What volume of H_2 forms at STP?

$n(Mg) = 0.120 / 24.3 = 4.94 \times 10^{-3}$ mol. $n(HCl) = 0.300 \times (20.0/1000) = 6.00 \times 10^{-3}$ mol.

Divide by coefficients: Mg $4.94\times10^{-3}/1 = 4.94\times10^{-3}$; HCl $6.00\times10^{-3}/2 = 3.00\times10^{-3}$. **HCl is limiting**.

$n(H_2) = 6.00\times10^{-3} / 2 = 3.00\times10^{-3}$ mol.

$V = 3.00\times10^{-3} \times 22.7 = 0.0681\ dm^3 =$ **68.1 cm^3**.

7. Multi-step and back-titration calculations (HL)

At HL the same tools are combined into longer chains: a back titration with several conversions, or two reactions run in sequence where the product of the first is the reactant of the second. Work in moles throughout and only convert to mass or volume at the very end.

Worked example 12 (HL) — back titration of impure limestone

A 2.50 g sample of crushed eggshell (mostly $CaCO_3$) is added to 50.0 cm^3 of 1.00 mol dm^{-3} HCl (an excess). The excess acid needs 24.0 cm^3 of 0.500 mol dm^{-3} NaOH to neutralise. Find the percentage of $CaCO_3$ in the shell.

Reactions: $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$; $HCl + NaOH \rightarrow NaCl + H_2O$.

Total HCl added = 1.00 × (50.0/1000) = 0.0500 mol.

Excess HCl = n(NaOH) = 0.500 × (24.0/1000) = 0.0120 mol.

HCl that reacted with the shell = 0.0500 − 0.0120 = 0.0380 mol.

$n(CaCO_3)$ = 0.0380 / 2 = 0.0190 mol; mass = 0.0190 × 100.1 = 1.902 g.

% $CaCO_3$ = (1.902 / 2.50) × 100 = **76.1%**.

Worked example 13 (HL) — two reactions in sequence

Ammonia is made and then catalytically oxidised (the first stage of nitric acid manufacture): $N_2 + 3H_2 \rightarrow 2NH_3$, then $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$. Starting from 84 g of N_2 with excess H_2 and O_2, and assuming complete conversion, find the mass of NO. (M: N_2 28.0, NO 30.0)

$n(N_2)$ = 84 / 28.0 = 3.0 mol → $n(NH_3)$ = 3.0 × 2 = 6.0 mol.

Ratio NH_3 : NO = 4 : 4 = 1 : 1, so n(NO) = 6.0 mol.

mass(NO) = 6.0 × 30.0 = **180 g**.

HL tip: In a long chain, keep unrounded intermediate values in your calculator and round only the final answer. Rounding at each step accumulates error and can cost the final accuracy mark.

8. Common pitfalls

- Balancing by changing a subscript instead of a coefficient → this changes the substance itself.
- Choosing the limiting reactant from mass or moles alone, without dividing by the coefficient.
- Leaving volumes in cm^3 in $n = c \times V$ → always convert to dm^3 first.
- Applying the mole ratio to masses or volumes directly instead of to moles.
- Forgetting the 2 in the ratio for diprotic acids (H_2SO_4) or dibasic hydroxides ($Ca(OH)_2$).
- Confusing percentage yield (depends on conditions) with atom economy (fixed by the equation).
- Using 24.0 dm^3 mol^{-1} (old room-temperature value); the 2025 booklet uses 22.7 dm^3 mol^{-1} at STP.
- Rounding intermediate answers too early in multi-step problems.

9. Quick reference

Quantity	Relationship
Moles from mass	$n = m / M$
Moles in solution	$n = c \times V$ (V in dm^3)
Moles of gas (STP)	$n = V / 22.7$ (V in dm^3)
$cm^3 \rightarrow dm^3$	divide by 1000
Percentage yield	(actual / theoretical) × 100%
Atom economy	(M_r desired product / M_r all products) × 100%
Limiting reactant	smallest value of (moles / coefficient)
Back titration	reacted = initial excess reagent − titrated excess

10. Test yourself

Attempt all ten without notes, then check against the full worked answers.

1. Balance: $Fe + O_2 \rightarrow Fe_2O_3$.

2. Write the net ionic equation for barium chloride solution mixed with sodium sulfate solution.

3. How many moles of O_2 are needed to burn 0.25 mol of butane? $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$.

4. 0.20 mol N_2 is mixed with 0.50 mol H_2. Identify the limiting reactant, the moles of NH_3, and the excess remaining.

5. 8.00 g NaOH reacts with 4.90 g H_2SO_4: $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$. Which is limiting, and what mass of Na_2SO_4 forms? (M: NaOH 40.0, H_2SO_4 98.1, Na_2SO_4 142.0)

6. 10.0 g of N_2 with excess H_2 gives 8.50 g of NH_3. Find the percentage yield. (M: N_2 28.0, NH_3 17.0)

7. Iron is extracted by $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$. Find the atom economy for iron. (M: Fe 55.8, CO_2 44.0)

8. 20.0 cm^3 of H_2SO_4 is neutralised by 25.0 cm^3 of 0.200 mol dm^{-3} KOH. $H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$. Find $[H_2SO_4]$.

9. What volume of CO_2 at STP forms when 5.00 g of $CaCO_3$ decomposes fully? (M: $CaCO_3$ 100.1)

10. (HL) A 2.00 g sample of impure Na_2CO_3 is dissolved and reacted with 50.0 cm^3 of 1.00 mol dm^{-3} HCl (excess). $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$. The excess HCl needs 32.0 cm^3 of 0.500 mol dm^{-3} NaOH. Find the % of Na_2CO_3. (M: Na_2CO_3 106.0)

Answers

1. $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ (Fe 4=4, O 6=6).

2. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$; Na^+ and Cl^- are spectators.

3. Ratio $C_4H_{10} : O_2 = 2 : 13$, so $n(O_2) = 0.25 \times (13/2) = 1.6$ mol.

4. Divide by coefficients: N_2 0.20/1 = 0.20; H_2 0.50/3 = 0.167. H_2 is limiting. $n(NH_3) = 0.50 \times (2/3) = 0.33$ mol. N_2 used = $0.50 \times (1/3) = 0.167$ mol, so N_2 excess = $0.20 - 0.167 = 0.033$ mol.

5. n(NaOH) = 8.00/40.0 = 0.200 mol; $n(H_2SO_4)$ = 4.90/98.1 = 0.0500 mol. Divide by coefficients: NaOH 0.200/2 = 0.100; acid 0.0500/1 = 0.0500. H_2SO_4 is limiting. $n(Na_2SO_4)$ = 0.0500 mol; mass = 0.0500 × 142.0 = 7.10 g.

6. $n(N_2)$ = 10.0/28.0 = 0.357 mol → NH_3 theoretical = 0.714 mol × 17.0 = 12.14 g. % yield = (8.50/12.14) × 100 = 70.0%.

7. Products per equation: 2 Fe (2×55.8 = 111.6) and 3 CO_2 (3×44.0 = 132.0). Atom economy = 111.6 / (111.6 + 132.0) × 100 = 45.8%.

8. n(KOH) = 0.200 × (25.0/1000) = 5.00 $\times 10^{-3}$ mol. Ratio KOH : H_2SO_4 = 2 : 1, so n(acid) = 2.50 $\times 10^{-3}$ mol. $[H_2SO_4]$ = 2.50 $\times 10^{-3}$ / (20.0/1000) = 0.125 mol dm^{-3}.

9. $n(CaCO_3)$ = 5.00/100.1 = 0.04995 mol = $n(CO_2)$. V = 0.04995 × 22.7 = 1.13 dm^3 (1130 cm^3).

10. Total HCl = 1.00 × (50.0/1000) = 0.0500 mol. Excess HCl = n(NaOH) = 0.500 × (32.0/1000) = 0.0160 mol. Reacted with Na_2CO_3 = 0.0500 − 0.0160 = 0.0340 mol. $n(Na_2CO_3)$ = 0.0340/2 = 0.0170 mol; mass = 0.0170 × 106.0 = 1.802 g. % = (1.802/2.00) × 100 = 90.1%.