

3 · Quantitative chemistry

AQA · GCSE · Chemistry

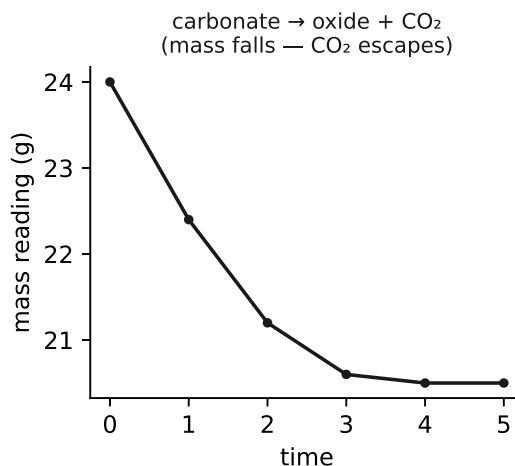
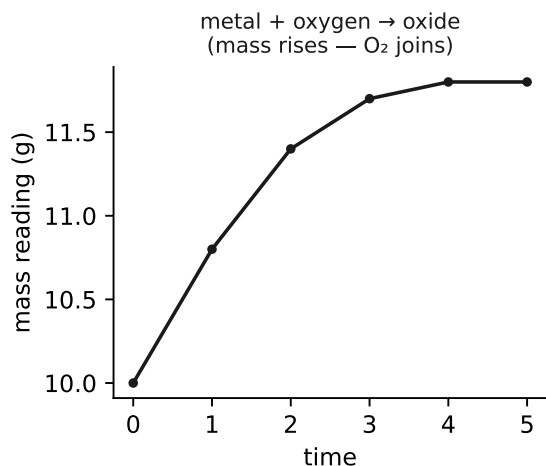
Quantitative chemistry: counting atoms by mass

- Atoms are neither lost nor made: the balanced equation is a mass ledger, and conservation of mass closes every account.
- The mole converts grams into particle counts; moles turn equations into reacting-mass arithmetic.
- Yield, atom economy, molar concentration and gas volumes (chemistry only, with the specified HT demands) finish the quantitative toolkit.

Conservation of mass and equations (4.3.1)

Conservation of mass : no atoms are lost or made, so the mass of products equals the mass of reactants. Symbol equations are **balanced** in atom numbers; know the difference between a **multiplier before a formula** (numbers of units) and a **subscript within it** (atoms in the unit).

Relative formula mass (Mr) = sum of the relative atomic masses in the numbers shown. In a balanced equation, multiply every relative formula mass by its equation coefficient before summing: the totals for reactants and products are equal. **Percentage by mass of an element** = $(Ar \times \text{number of atoms} \div Mr) \times 100 \%$.



Mass changes with gases: metal + oxygen → oxide **gains** mass; thermal decomposition of a carbonate **loses** the escaped CO₂ – no law is broken once the gas is counted.

Uncertainty: every measurement carries uncertainty; use the **range of repeated measurements about the mean** to estimate the spread. A mean $\pm$ half-range

report is a stated classroom convention, not a rule that automatically includes every source of measurement error.

■ Count atoms before comparing masses

For $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$, the left contains two Mg atoms and two O atoms; so does the right. With $\text{Mg} = 24$ and $\text{O} = 16$, the relative mass total is $2(24) + 32 = 80$ on the left and $2(40) = 80$ on the right. The unweighted calculation $24 + 32 = 40$ is false. Changing the coefficient changes how many formula units react; changing a subscript changes the substance.

Worked formula calculation. For $(\text{NH}_4)_2\text{SO}_4$, count $\text{N}_2\text{H}_8\text{SO}_4$: $M_r = 2(14) + 8(1) + 32 + 4(16) = 132$. Sulfur's mass fraction is $32/132$, so its percentage by mass is $32/132 \times 100 = 24.2\%$. Use the mass of the chosen element throughout the whole formula, including bracket multipliers. Percentages for all elements add to approximately 100%, allowing rounding.

Worked gas accounting. Heating 10.0 g CaCO_3 produces 5.6 g CaO while 4.4 g CO_2 escapes. The balance under the open vessel falls by 4.4 g; the full reaction still accounts for $5.6 + 4.4 = 10.0$ g. During oxidation, oxygen enters from outside the weighed metal sample, so the solid can gain mass. Explain which particles crossed the system boundary. Never seal a heated gas-producing reaction to make a classroom mass demonstration: use an approved apparatus and account for the gas.

Worked uncertainty report. Repeated masses are 2.40, 2.42 and 2.44 g. Mean = $(2.40 + 2.42 + 2.44)/3 = 2.42$ g; range = $2.44 - 2.40 = 0.04$ g. Under the stated half-range convention report 2.42 ± 0.02 g. Plot individual readings on a number line to show their distribution. A close cluster does not prove the balance is correctly calibrated. Retain unusual readings unless there is an evidenced reason to exclude them; report any exclusion and reason.

Moles and reacting masses – HT (4.3.2)

The mole : the mass of one mole of a substance in grams is numerically its **Mr**. One mole contains the **Avogadro constant** of particles – 6.02×10^{23} – the same count of stated particles (atoms, molecules, ions) as a mole of any other substance.

Use the equation and its units directly. $n = m/M$; rearranging gives $m = nM$ and $M = m/n$, where molar mass M has units g/mol and is numerically equal to the relative formula mass. Relative formula mass itself has no unit.

$$\text{moles} = \frac{\text{mass (g)}}{M_r}$$

Equations as mole ratios: $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$ reads “1 mol Mg reacts with 2 mol HCl $\rightarrow$ 1 mol MgCl_2 + 1 mol H_2 ”. Given any one mass, calculate all others: mass $\rightarrow$ moles $\rightarrow$ ratio $\rightarrow$ moles $\rightarrow$ mass.

Balancing from masses: convert each mass to moles, divide by the smallest, clear fractions to whole numbers.

Limiting reactant : the reactant completely used up limits the product; an **excess** of the other ensures completion. Calculate the product from the limiting reactant's moles only.

Concentration : in g/dm^3 for all tiers – mass of solute in a given volume; (HT) also mol/dm^3 : $\text{moles} = \text{concentration} \times \text{volume}(\text{dm}^3)$, rearranged as needed.

■ Choose the particle and the equation ratio

Worked mole conversion (HT). For 8.0 g NaOH, $M = 23 + 16 + 1 = 40 \text{ g/mol}$, so $n = 8.0 \text{ g}/(40 \text{ g/mol}) = 0.20 \text{ mol}$. Formula-unit count $= 0.20 \times 6.02 \times 10^{23} = 1.20 \times 10^{23}$. Each formula unit contains one O and one H, so it contains twice as many O-plus-H atoms as formula units. Always name what is counted: a mole of O_2 molecules contains two moles of oxygen atoms. The Avogadro constant applies to stated ions and electrons as well as atoms and molecules.

Worked reacting mass (HT). Find the CO_2 from complete decomposition of 10.0 g CaCO_3 . $M(\text{CaCO}_3) = 100 \text{ g/mol}$, so $n = 10.0/100 = 0.100 \text{ mol}$. The equation $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ has a 1:1 ratio, giving 0.100 mol CO_2 . Its molar mass is 44 g/mol, so $m = 0.100 \times 44 = 4.40 \text{ g}$. Reverse the chain if the product mass is supplied: product mass $\rightarrow$ product moles $\rightarrow$ reactant moles $\rightarrow$ reactant mass. Mass ratios are not generally the same as mole ratios.

Worked balancing from masses (HT). Suppose 4.8 g Mg combines with 3.2 g O_2 to form 8.0 g MgO. Amounts are $4.8/24 = 0.20 \text{ mol Mg}$, $3.2/32 = 0.10 \text{ mol O}_2$ and $8.0/40 = 0.20 \text{ mol MgO}$. Divide all by 0.10 to obtain 2:1:2. Write $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$ and verify atoms and mass. Use the formula of the actual oxygen reactant, O_2 ; dividing its mass by 16 would count oxygen atoms instead of oxygen molecules.

Worked limiting reactant (HT). Mix 4.8 g Mg with 6.4 g O_2 . Amounts are 0.20 mol Mg and 0.20 mol O_2 . Compare amount divided by coefficient: Mg $0.20/2 = 0.10$; O_2 $0.20/1 = 0.20$. Magnesium runs out first. It makes 0.20 mol MgO, mass $= 0.20 \times 40 = 8.0 \text{ g}$. Oxygen used $= 0.20/2 = 0.10 \text{ mol}$, leaving 0.10 mol $= 3.2 \text{ g}$. Check total mass: $4.8 + 6.4 = 8.0 + 3.2 = 11.2 \text{ g}$. More oxygen cannot increase the product once magnesium is exhausted. Exactly stoichiometric amounts can use up both reactants together; do not force a single limiting label in that case.

Worked mass concentration (both tiers). A 250 cm^3 solution contains 3.5 g solute. Convert volume: $250/1000 = 0.250 \text{ dm}^3$. Concentration $= 3.5/0.250 = 14 \text{ g/dm}^3$. A 100 cm^3 portion contains $m = cV = 14 \times 0.100 = 1.4 \text{ g}$ solute. Volume means the final solution volume, not automatically the volume of water added. At fixed solute mass, doubling volume halves concentration; at fixed volume, doubling solute mass doubles concentration (HT explanation).

Yield and atom economy – chemistry only (4.3.3)

Percentage yield is below 100 % because: the reaction is **reversible**; product is **lost on separation**; reactants react in **unwanted ways**.

$$\% \text{ yield} = \frac{\text{mass actually made}}{\text{maximum theoretical mass}} \times 100$$

Atom economy measures how much of the starting material ends up in the **useful product** – high atom economy matters for sustainability and cost:

$$\text{atom economy} = \frac{\text{coefficient} \times M_r \text{ of desired product}}{\text{sum of coefficient} \times M_r \text{ for all reactants}} \times 100 \%$$

(HT) calculate the theoretical mass from the balanced equation, then the yield.

■ Separate actual yield from equation efficiency

Worked yield. If the theoretical product mass is 8.0 g and the dry recovered product is 6.0 g, $\text{yield} = 6.0/8.0 \times 100 = 75\%$. A yield above 100% suggests the weighed material is wet, contaminated or affected by measurement/calculation error; it does not show extra atoms were created. When theoretical mass is not supplied, first calculate it from the balanced equation and limiting amount (HT), then compare the same product's actual and theoretical masses.

Worked atom economy. For $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ with CaO desired, $\text{atom economy} = 56/100 \times 100 = 56\%$. If CO_2 were instead the desired product, it would be 44%. For $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, desired-product relative mass is $2 \times 18 = 36$, and total reactant relative mass = $2 \times 2 + 32 = 36$, so atom economy is 100%. Count every coefficient; using only one H_2O would falsely give 50%. These values come from the equation, whereas actual yield comes from the experiment. A 100% atom-economy reaction can still give a low isolated yield.

Choose a route (HT). A fictional route A has 90% atom economy but 40% yield and requires prolonged heating. Route B has 75% atom economy, 95% yield and a useful saleable by-product. A judgement needs more than picking the largest first percentage: compare useful output, energy, speed, equilibrium/reversibility, waste treatment and by-product demand under the supplied conditions. State the tradeoff and missing evidence, rather than declaring one number proves sustainability.

Concentrations in mol/dm^3 – chemistry only, HT (4.3.4)

Concentration in mol/dm^3 links moles, mass and volume: $\text{moles} = C \times V$; from a titration, knowing the volumes of both solutions and one concentration gives the other – use the balanced equation to relate acid and alkali amounts. Equal volumes do not imply equal amounts, and acid moles equal alkali moles only for a 1:1 equation ratio.

■ Concentration is amount per final solution volume

Use $n = cV$ with V in dm^3 for c in mol/dm^3 . Rearranging gives $c = n/V$ or $V = n/c$. A solution at 0.20 mol/dm^3 contains $0.20 \text{ mol per dm}^3$; it does not contain 0.20 mol in every beaker regardless of volume. To convert molar concentration to mass concentration, multiply by molar mass: $0.20 \text{ mol/dm}^3 \text{ NaOH} \times 40 \text{ g/mol} = 8.0 \text{ g/dm}^3$.

Worked non-1:1 titration. $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$. A 25.0 cm^3 acid sample reacts completely with 20.0 cm^3 of $0.100 \text{ mol/dm}^3 \text{ NaOH}$. Alkali amount = $0.100 \times 0.0200 = 0.00200 \text{ mol}$. Acid amount = $0.00200/2 = 0.00100 \text{ mol}$. Acid concentration = $0.00100/0.0250 = 0.0400 \text{ mol/dm}^3$. The volumes are both known, but the equation determines the mole ratio. Check by multiplying acid amount by two to recover the required alkali amount. Titration apparatus and method are developed in Chemical changes; this example interprets completed-reaction data.

Gas volumes – chemistry only, HT (4.3.5)

A given volume of gas contains the same number of moles at the same temperature and pressure – **equal volumes = equal moles**. The volumes of reacting gases (and products) follow the equation's ratio directly, e.g. 2 volumes of hydrogen react with 1 volume of oxygen.

■ Use matching temperature, pressure and gas units

At room temperature and pressure (20°C and one atmosphere), one mole of gas occupies 24 dm^3 in this specification's approximation. Thus $V = n \times 24 \text{ dm}^3$ and $n = V/24$ when V is in dm^3 . Convert cm^3 to dm^3 first. This value does not apply unchanged at any temperature or pressure, and a volume ratio applies to gases measured under the same conditions.

Worked mass to gas volume. For 4.4 g CO_2 at room conditions, $n = 4.4/44 = 0.100 \text{ mol}$, so $V = 0.100 \times 24 = 2.40 \text{ dm}^3 = 2400 \text{ cm}^3$. The reverse is $m = (V/24)M$ with consistent volume units.

Worked gas reaction. $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$. At identical temperature and pressure, $60 \text{ cm}^3 \text{ N}_2$ requires $180 \text{ cm}^3 \text{ H}_2$ and can form $120 \text{ cm}^3 \text{ NH}_3$ if reaction is complete. With only $150 \text{ cm}^3 \text{ H}_2$, hydrogen limits the result: N_2 used = $150/3 = 50 \text{ cm}^3$, NH_3 formed = $150 \times 2/3 = 100 \text{ cm}^3$ and N_2 remaining = $60 - 50 = 10 \text{ cm}^3$. A real reversible reaction may not reach that theoretical conversion. Do not assign the gas volume ratio to a solid or condensed liquid product merely because it has a coefficient in the equation.

Checklist before you call this topic done

- Balance atoms and weighted masses; explain gas mass changes and measurement uncertainty.
- Apply the relevant tier's mole, yield, atom-economy, concentration and gas calculations with units and conditions.