How much? The amount of chemical change

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

This handout teaches Reactivity 2.1, How much? The amount of chemical change, for the course first assessed in 2025. All of it is SL and HL content.

Prerequisites: the mole, molar mass and concentration from Structure 1.4, and the molar volume of a gas from Structure 1.5. Use the relative atomic masses from the data booklet.

Balanced equations

A recipe for a reaction

A cake recipe says 2 eggs for every 250 g of flour. With 6 eggs you need 750 g of flour, and if you have only 500 g of flour, the flour runs out first. A balanced equation is a recipe for a reaction, but it counts particles, not eggs.

Conservation of mass

In a chemical reaction atoms are rearranged, never made or destroyed. This is the **conservation of mass** : the total mass of the products equals the total mass of the reactants. So a balanced equation has the same number of atoms of each element on both sides.

Balancing an equation

Change only the numbers in front of the formulae, the **coefficients** . Never change a formula itself.

Propane burns in oxygen:

- Unbalanced: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$.
- Balance C: 3 C on the left, so $3CO_2$.
- Balance H: 8 H on the left, so $4H_2O$.
- Balance O last: $3 \times 2 + 4 \times 1 = 10$ O atoms on the right, so $5O_2$.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

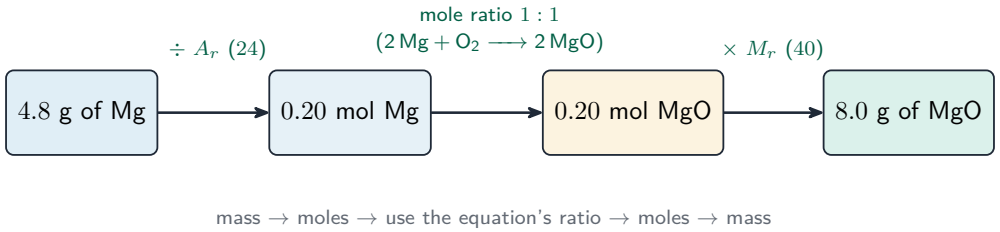
State symbols: (s) solid, (l) liquid, (g) gas, (aq) dissolved in water.

What the coefficients mean

The coefficients give the **mole ratio** of the reaction. For propane: 1 mol of C_3H_8 reacts with 5 mol of O_2 to give 3 mol of CO_2 and 4 mol of H_2O. The study of these amounts is **stoichiometry** .

Reacting masses

Three steps



Mass → moles → use the equation's ratio → moles → mass

1. Convert the mass you know to moles: $n = m/M$.
2. Use the mole ratio from the balanced equation.
3. Convert the moles you want back to a mass: $m = n \times M$.

Worked example: lime from limestone

What mass of calcium oxide forms when 50.0 g of calcium carbonate is heated?

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

- $n(CaCO_3) = 50.0/100.09 = 0.500$ mol.
- Ratio 1 : 1, so $n(CaO) = 0.500$ mol.
- $m(CaO) = 0.500 \times 56.08 = 28.0$ g.

Worked example: iron from iron ore

What mass of iron forms from 16.0 g of iron(III) oxide?

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$$

- $n(Fe_2O_3) = 16.0/159.70 = 0.100$ mol.
- Ratio Fe_2O_3 : Fe = 1 : 2, so $n(Fe) = 0.200$ mol.
- $m(Fe) = 0.200 \times 55.85 = 11.2$ g.

Reacting gas volumes

Using the molar volume

At STP (273 K, 100 kPa) one mole of any ideal gas occupies 22.7 dm^3. This is the **molar volume** (data booklet).

$$n = \frac{V}{22.7} \quad (V \text{ in } dm^3 \text{ at STP})$$

Worked example: gas from a reaction

What volume of carbon dioxide, at STP, forms when 2.50 g of calcium carbonate reacts with excess hydrochloric acid?

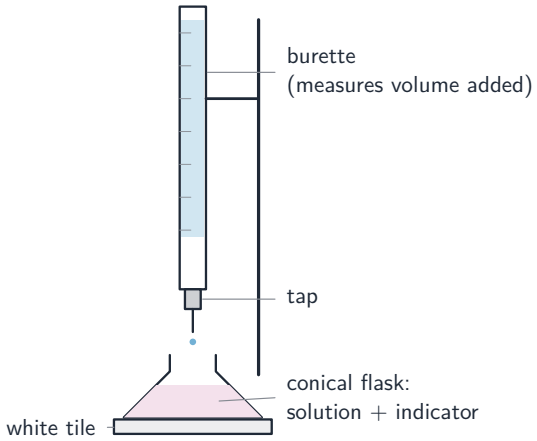
$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

- $n(CaCO_3) = 2.50/100.09 = 0.0250$ mol.
- Ratio 1 : 1, so $n(CO_2) = 0.0250$ mol.
- $V = 0.0250 \times 22.7 = 0.567\ dm^3 = 567\ cm^3$.

For two gases at the same temperature and pressure you do not need the molar volume: the volume ratio is the mole ratio (Avogadro's law, Structure 1.4).

Titrations

What a titration does



The titre is the final burette reading minus the initial reading

A **titration** finds the concentration of a solution by reacting it with a **standard solution** of known concentration.

- A **pipette** delivers an exact volume, often 25.00 cm^3, of one solution into a conical flask.
- Add a few drops of **indicator** .
- Add the other solution from a **burette** until the indicator just changes colour: the **end point** .
- The volume added from the burette is the **titre** .

Good technique

- Rinse the burette with the solution it will hold, and the pipette with the solution it will deliver. Rinse the flask with distilled water only; extra water does not change the amount in it.
- Stand the flask on a white tile so the colour change is easy to see.
- Do a rough titration first. Then add the solution drop by drop near the end point, swirling the flask.
- Read the bottom of the meniscus at eye level, to the nearest 0.05 cm^3.
- Repeat until two or more titres are **concordant** , within 0.10 cm^3 of each other, and average only those.

Worked example: a titration calculation

25.00 cm^3 of sodium hydroxide solution is titrated with 0.100 mol dm^{-3} sulfuric acid.

Titration	**rough**	**1**	**2**	**3**
Final reading / cm^3	22.10	43.65	21.45	42.95
Initial reading / cm^3	0.00	22.10	0.00	21.45
Titre / cm^3	22.10	21.55	21.45	21.50

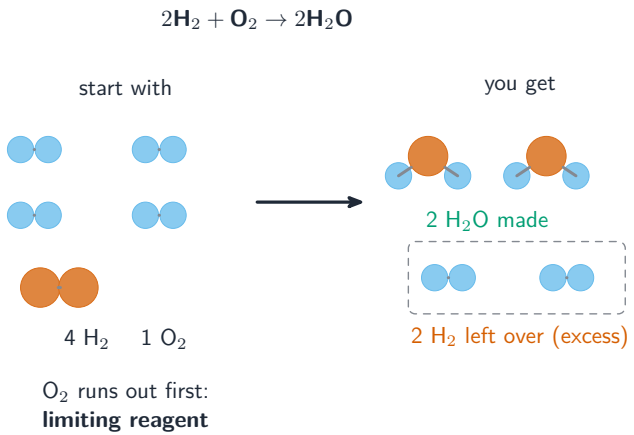
The concordant titres are 21.55, 21.45 and 21.50 cm^3; the mean is 21.50 cm^3. The rough titre is left out.

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

- $n(H_2SO_4) = 0.100 \times 21.50/1000 = 2.150 \times 10^{-3}$ mol.
- Ratio 1 : 2, so $n(NaOH) = 4.300 \times 10^{-3}$ mol.
- $c(NaOH) = 4.300 \times 10^{-3}/0.02500 = 0.172$ mol dm^{-3}.

Limiting reactant and yield

The limiting reactant



The reactant that runs out first decides how much product forms

The **limiting reactant** is used up first and stops the reaction. The other reactant is **in excess**: some of it is left over.

To find the limiting reactant, work out the amount of each reactant in moles and **divide by its coefficient** in the equation. The smaller result is the limiting reactant.

Worked example: aluminium and chlorine

5.40 g of aluminium is heated with 14.2 g of chlorine.

$$2\mathrm{Al(s)} + 3\mathrm{Cl_2(g)} \rightarrow 2\mathrm{AlCl_3(s)}$$

- $n(\mathrm{Al}) = 5.40/26.98 = 0.200$ mol; $0.200/2 = 0.100$.
- $n(\mathrm{Cl_2}) = 14.2/70.90 = 0.200$ mol; $0.200/3 = 0.0667$.
- 0.0667 is smaller, so chlorine is the limiting reactant and aluminium is in excess.
- $n(\mathrm{AlCl_3}) = 0.200 \times 2/3 = 0.133$ mol; $m = 0.133 \times 133.33 = 17.8$ g.

Equal amounts in moles do not mean neither is limiting. Always use the coefficients.

Percentage yield

The **theoretical yield** is the mass of product calculated from the limiting reactant. The mass you actually collect is the actual yield.

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

If 14.2 g of $\mathrm{AlCl_3}$ is collected in the example above: $14.2/17.8 \times 100 = 79.8\%$.

Yields are less than 100% because:

- the reaction may be reversible and reach equilibrium;
- side reactions make other products;
- product is lost when it is filtered, transferred or purified.

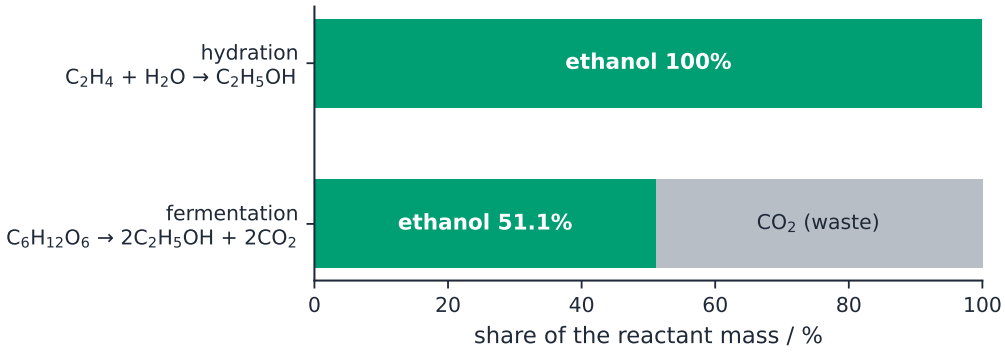
Atom economy

How much of the reactants ends up as product?

Atom economy measures how much of the mass of the reactants ends up in the product you want.

$$\text{atom economy} = \frac{\text{molar mass of the desired product}}{\text{total molar mass of all reactants}} \times 100$$

Use the coefficients from the balanced equation in both totals.



An addition reaction has no waste product, so its atom economy is 100%

Worked example: two routes to ethanol

- Hydration: $\mathrm{C_2H_4 + H_2O \rightarrow C_2H_5OH}$. Every atom ends up in ethanol: atom economy 100%.
- Fermentation: $\mathrm{C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2}$. Atom economy $= 2 \times 46.08/180.18 \times 100 = 51.1\%$.

A high atom economy means less waste, less raw material and often less cost; it is one idea of green chemistry. It is not the same as percentage yield: a reaction can have a 100% atom economy but a poor yield.

Practice

1. Balance: $\mathrm{Al + O_2 \rightarrow Al_2O_3}$.
2. Balance: $\mathrm{C_2H_6 + O_2 \rightarrow CO_2 + H_2O}$.
3. Calculate the mass of magnesium oxide made by burning 6.08 g of magnesium: $\mathrm{2Mg + O_2 \rightarrow 2MgO}$.
4. Calculate the mass of oxygen needed to burn 11.0 g of propane completely (equation in the handout).

5. Calculate the volume of hydrogen at STP made when 0.654 g of zinc reacts with excess acid: $\mathrm{Zn} + 2\mathrm{HCl} \rightarrow \mathrm{ZnCl_2} + \mathrm{H_2}$.
6. 25.00 $\mathrm{cm^3}$ of hydrochloric acid needs 18.40 $\mathrm{cm^3}$ of 0.150 mol $\mathrm{dm^{-3}}$ sodium hydroxide for neutralization. Calculate the concentration of the acid.
7. Titres of 24.30, 23.80, 23.70 and 23.75 $\mathrm{cm^3}$ were recorded; the first was the rough titration. Calculate the mean titre to use.
8. Explain why the conical flask is rinsed with distilled water but the burette is rinsed with the solution it holds.
9. 4.00 g of hydrogen reacts with 16.0 g of oxygen: $2\mathrm{H_2} + \mathrm{O_2} \rightarrow 2\mathrm{H_2O}$. Find the limiting reactant and the mass of water made.
10. 10.0 g of $\mathrm{CaCO_3}$ is heated and 4.80 g of CaO is collected. Calculate the percentage yield.
11. Calculate the atom economy for making hydrogen from methane: $\mathrm{CH_4} + \mathrm{H_2O} \rightarrow \mathrm{CO} + 3\mathrm{H_2}$.
12. Give two reasons why the actual yield of a reaction is often less than the theoretical yield.

Answers.

1. $4\mathrm{Al} + 3\mathrm{O_2} \rightarrow 2\mathrm{Al_2O_3}$.
2. $2\mathrm{C_2H_6} + 7\mathrm{O_2} \rightarrow 4\mathrm{CO_2} + 6\mathrm{H_2O}$.
3. $n(\mathrm{Mg}) = 6.08/24.31 = 0.250$ mol, so 0.250 mol MgO; $m = 0.250 \times 40.31 = 10.1$ g.
4. $n(\mathrm{C_3H_8}) = 11.0/44.11 = 0.249$ mol; $n(\mathrm{O_2}) = 5 \times 0.249 = 1.25$ mol; $m = 1.25 \times 32.00 = 39.9$ g.
5. $n(\mathrm{Zn}) = 0.654/65.38 = 0.0100$ mol, so 0.0100 mol $\mathrm{H_2}$; $V = 0.0100 \times 22.7 = 0.227$ $\mathrm{dm^3}$ (227 $\mathrm{cm^3}$).
6. $n(\mathrm{NaOH}) = 0.150 \times 0.01840 = 2.76 \times 10^{-3}$ mol; ratio 1 : 1; $c(\mathrm{HCl}) = 2.76 \times 10^{-3}/0.02500 = 0.110$ mol $\mathrm{dm^{-3}}$.
7. Concordant titres 23.80, 23.70, 23.75; mean 23.75 $\mathrm{cm^3}$.
8. Water in the flask does not change the amount of solute it holds. Water left in the burette would dilute the solution and make the titre too large.
9. $n(\mathrm{H_2}) = 4.00/2.02 = 1.98$ mol, $\div 2 = 0.990$; $n(\mathrm{O_2}) = 16.0/32.00 = 0.500$ mol, $\div 1 = 0.500$. Oxygen limits. $n(\mathrm{H_2O}) = 1.00$ mol; $m = 1.00 \times 18.02 = 18.0$ g.
10. $n(\mathrm{CaCO_3}) = 10.0/100.09 = 0.0999$ mol; theoretical CaO $= 0.0999 \times 56.08 = 5.60$ g; yield $= 4.80/5.60 \times 100 = 85.7\%$.
11. Reactants $16.05 + 18.02 = 34.07$; desired product $3 \times 2.02 = 6.06$; atom economy $= 6.06/34.07 \times 100 = 17.8\%$.
12. Any two: the reaction is reversible; side reactions; product lost in transfer, filtering or purification; the reaction does not go to completion.

Exam tips

- Balance the equation before any calculation, and change only the coefficients.
- Show the mole ratio as a separate line; it is where most marks are lost.
- Divide $\mathrm{cm^3}$ by 1000 before using $c = n/V$, and use 22.7 $\mathrm{dm^3}$ only at STP.
- Average only concordant titres, and leave out the rough titration.
- For the limiting reactant, divide moles by the coefficient; never compare masses.
- Percentage yield compares masses of product; atom economy compares molar masses in the equation.