

Stoichiometry

IGCSE Chemistry 0620 · Topic 3 revision notes

Everything from the 6 lessons of this topic on a few pages: the key idea of each lesson, the diagram to remember, the points to learn and the traps examiners set. Read it, cover it, test yourself.

1. Formulae and equations

2. Relative masses

3. The mole and gas volumes

4. Reacting masses, yield and purity

5. Concentration and titrations

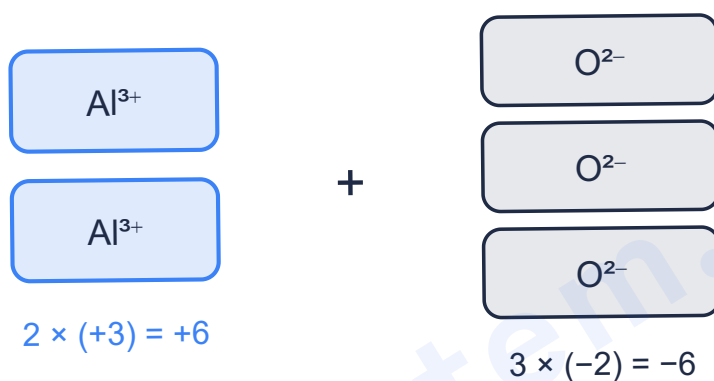
6. Empirical and molecular formulae

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1 Formulae and equations

Syllabus 3.1

In an ionic compound the total positive charge equals the total negative charge, so Al^{3+} and O^{2-} combine 2 : 3 as Al_2O_3 . A balanced symbol equation has the same number of each type of atom on both sides. Balance it by changing only the big numbers in front of formulae, never the small numbers inside them. State symbols show the state of each substance: (s) solid, (l) liquid, (g) gas, (aq) dissolved in water.



total charge 0, so the formula is Al_2O_3

Fig. 1: the charges cancel to fix the formula of aluminium oxide

Formula first, balance second.

- ✓ **Ionic formula:** total positive charge = total negative charge.
- ✓ **Brackets** round a group such as OH^- or NH_4^+ when there is more than one.
- ✓ **Balance with the numbers in front only;** count O atoms last in combustion.
- ✓ **Ionic equation (Extended):** leave out spectator ions; atoms and charges must both balance.

Watch out!

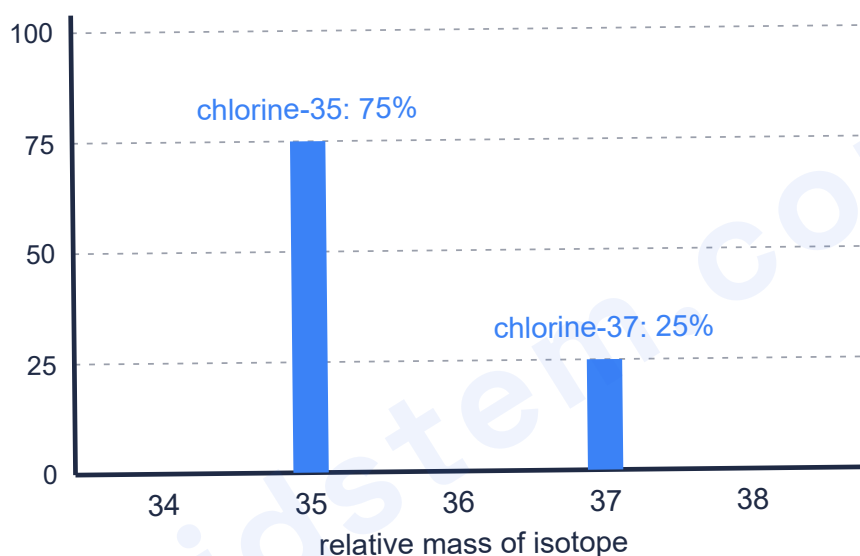
Never balance an equation by changing a small number inside a formula. Writing H_2O_2 to make $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$ balance turns water into hydrogen peroxide, a different substance. Change only the numbers in front: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$.

2 Relative masses

Syllabus 3.2

Relative atomic mass A_r is the average mass of the isotopes of an element compared with 1/12 of the mass of an atom of carbon-12. Relative molecular mass M_r is the sum of the A_r values of all the atoms in a formula; the same sum is used for ionic compounds (relative formula mass). Relative masses are ratios, so they have no units.

abundance / %



$$A_r = (35 \times 75 + 37 \times 25) \div 100 = 35.5$$

Fig. 1: the two isotopes of chlorine and the average they give

Everything is compared with carbon-12.

- ✓ **Standard:** 1/12 of the mass of an atom of carbon-12.
- ✓ **M_r :** add the A_r of every atom; a number after a bracket multiplies everything inside.
- ✓ **% by mass** = $(A_r \times \text{number of atoms}) \div M_r \times 100$.
- ✓ **No units:** relative masses are ratios.

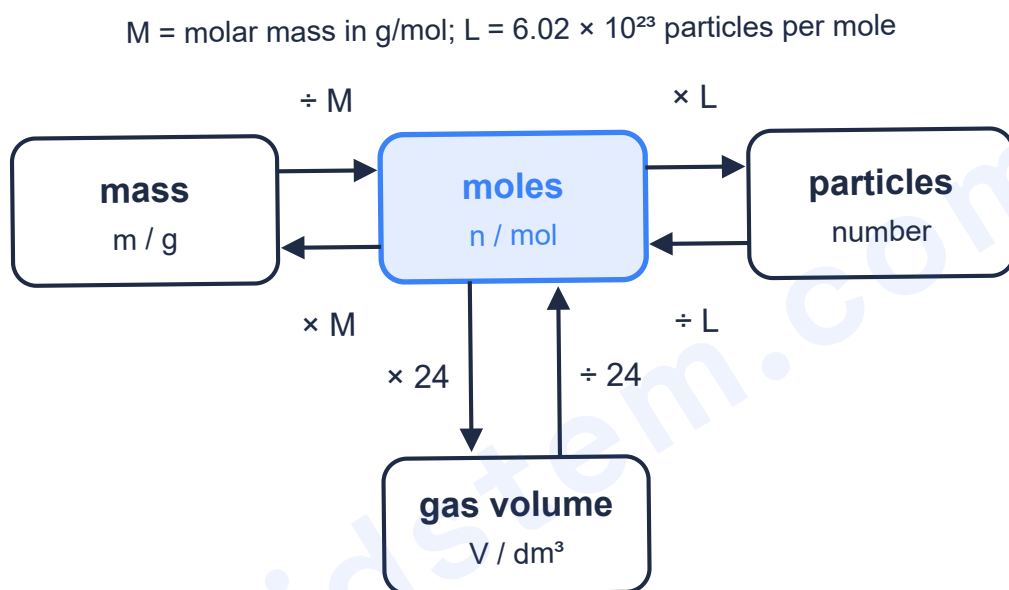
Watch out!

Count **every** atom of the element. Ammonium nitrate, NH_4NO_3 , has one N in the ammonium ion and one in the nitrate ion, so the nitrogen mass is 28, not 14. Using 14 gives 17.5%, exactly half the right answer.

3 The mole and gas volumes

Syllabus 3.3

One mole of a substance contains 6.02×10^{23} particles (the Avogadro constant). The mass of one mole in grams equals its A_r or M_r , so $\text{moles} = \text{mass} \div M_r$. One mole of **any gas** occupies 24 dm^3 at room temperature and pressure (r.t.p.), so $\text{moles of gas} = \text{volume in dm}^3 \div 24$, or $\text{volume in cm}^3 \div 24\,000$.



gases only: 1 mol of any gas occupies 24 dm^3 at r.t.p.

Fig. 1: every route between mass, particles and gas volume goes through the mole

Go through the mole every time.

- ✓ $\text{moles} = \text{mass} \div M_r$ with the mass in grams.
- ✓ $\text{particles} = \text{moles} \times 6.02 \times 10^{23}$; say which particle: atoms or molecules.
- ✓ $\text{Gas moles} = \text{volume in dm}^3 \div 24$ (or $\text{cm}^3 \div 24\,000$) at r.t.p.
- ✓ **Gases only** for the 24 dm³ rule.

Watch out!

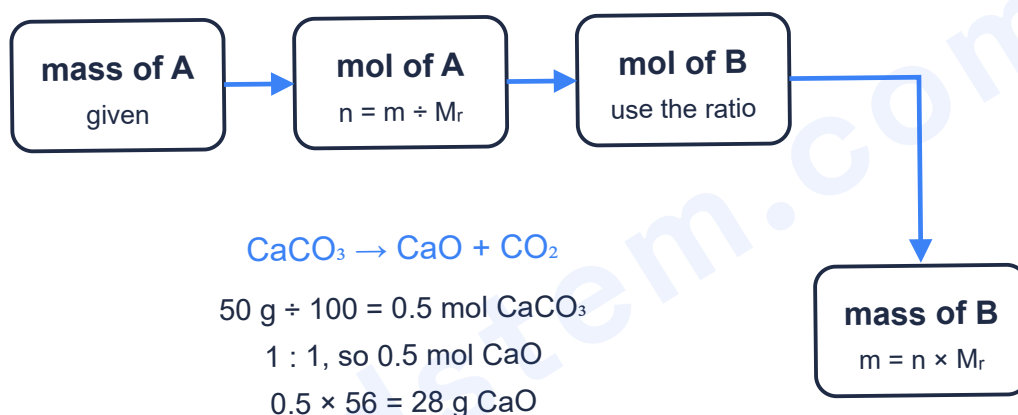
Read whether the question wants **molecules** or **atoms**. 9 g of water is 3.01×10^{23} molecules but 9.03×10^{23} atoms, because each H₂O molecule contains three atoms.

The 24 dm³ rule is for **gases only**, at r.t.p. Never use it for a solid, a liquid or a solution. And check the unit: dividing a volume in cm³ by 24 gives an answer 1000 times too big.

4 Reacting masses, yield and purity

Syllabus 3.3

Use three steps: (1) change the mass (or gas volume) you are given into moles; (2) use the numbers in the balanced equation to find the moles of the substance you want; (3) change those moles back into a mass or volume. The **limiting reactant** is the one that is completely used up; it fixes the amount of product. Percentage yield = $\text{actual mass} \div \text{expected mass} \times 100$. Percentage purity = $\text{mass of pure substance} \div \text{mass of impure sample} \times 100$.



never compare masses directly: the equation counts moles

Fig. 1: the three-step route from mass of A to mass of B

Moles, ratio, back again.

- ✓ **Step 1:** moles = mass $\div$ M_r (or gas volume $\div$ 24).
- ✓ **Step 2:** use the numbers in the balanced equation; M_r never includes them.
- ✓ **Step 3:** mass = moles $\times$ M_r (or volume = moles $\times$ 24).
- ✓ **Limiting reactant first,** then calculate the product from it.

Watch out!

The 2 in front of 2MgO belongs in the mole ratio, not in the M_r . Use $M_r(\text{MgO}) = 40$, not 80. A student who calculates $0.2 \text{ mol} \times 80$ gets 16 g, double the correct 8.0 g.

Equal numbers of moles do not mean nothing is left over. Here there are 0.1 mol of each, but the equation needs twice as much HCl as Mg, so HCl runs out first.

Always compare the moles you have with the ratio in the equation.

5 Concentration and titrations

Syllabus 3.3

Concentration can be given in g/dm^3 or in mol/dm^3 . Moles of solute = concentration in $\text{mol/dm}^3 \times \text{volume in dm}^3$, and to change cm^3 into dm^3 you divide by 1000. In a titration you find the moles of the solution you know completely, use the equation's ratio, and then divide by the volume of the other solution in dm^3 .

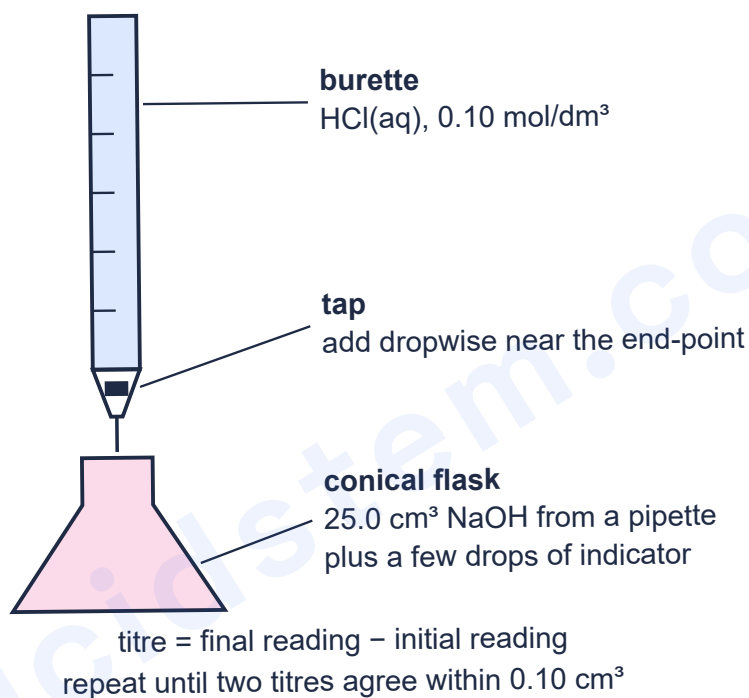


Fig. 1: titration apparatus for finding the concentration of an alkali

Known solution first.

- ✓ $\text{cm}^3 \div 1000 = \text{dm}^3$ before any calculation.
- ✓ $\text{moles} = \text{concentration} \times \text{volume in dm}^3$.
- ✓ Ratio from the balanced equation;
 H_2SO_4 reacts 1 : 2 with NaOH.
- ✓ $\text{g/dm}^3 = \text{mol/dm}^3 \times M_r$

Watch out!

Burettes and pipettes measure in cm^3 , but concentrations are per dm^3 . Forgetting to divide by 1000 makes 25 cm^3 of 0.2 mol/dm^3 look like 5 mol instead of 0.005 mol.

Do not skip the ratio. Treating the sulfuric acid titration as 1 : 1 gives $0.0025 \div 0.0125 = 0.20 \text{ mol/dm}^3$, double the right answer. Write the balanced equation first and read the ratio from it.

6 Empirical and molecular formulae

Syllabus 3.3

The **empirical formula** is the simplest whole-number ratio of the different atoms or ions in a compound. The **molecular formula** is the number and type of each atom in one molecule. To find the empirical formula, divide each mass (or percentage) by A_r , then divide by the smallest answer. To find the molecular formula, divide M_r by the empirical formula mass and multiply the empirical formula by the answer.

step	C	H	O
% by mass	40.0	6.7	53.3
$\div A_r$	$\div 12 = 3.33$	$\div 1 = 6.7$	$\div 16 = 3.33$
$\div$ smallest	1.00	2.01	1.00
ratio	1	2	1

empirical formula CH_2O (mass 30)

$M_r = 180$, so $180 \div 30 = 6$: molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$

Fig. 1: from percentages to glucose, step by step

Divide, divide, make whole.

- ✓ **Divide each mass or % by A_r** , then by the smallest answer.
- ✓ **Double** if a value ends in .5; never round it away.
- ✓ **Molecular formula:** $M_r \div$ empirical formula mass = the multiplier.
- ✓ **Find a missing mass** by subtraction: mass of oxide – mass of metal = mass of oxygen.

Watch out!

Never round 1 : 1.5 to 1 : 2 (or to 1 : 1). A ratio ending in .5 means you must double every number. Rounding the iron oxide ratio gives FeO_2 , which is wrong; doubling gives Fe_2O_3 . And never divide the masses themselves: 2.4 g Mg to 1.6 g O is a 1 : 1 ratio of atoms, not 3 : 2.