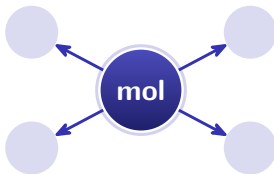


Stoichiometry

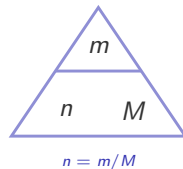
IGCSE Chemistry ■ Topic 3

IGCSE Chemistry



What we will cover

- 1 Formulae & equations
- 2 The mole & molar mass
- 3 Gas volumes & concentration
- 4 Reaction calculations
- 5 Limiting reactant & titration
- 6 Empirical formulae & percentages



the mole and molar mass



Chemical formulae



the charge numbers cross over to become the subscripts (then simplify if they share a factor)

- The **molecular formula** is the actual number of each atom in one **molecule** . For glucose it is $\text{C}_6\text{H}_{12}\text{O}_6$.
- The **empirical formula** is the simplest whole-number **ratio** of the atoms or **ions** in a **compound** . For glucose it is CH_2O .

Formulae, equations and state symbols

an ionic compound

work the formula out from the **charges**: total positive must balance total negative

the common ions

+: Na^+ , K^+ , Ca^{2+} , Mg^{2+} , Al^{3+} , NH_4^+ . -: Cl^- , OH^- , NO_3^- , SO_4^{2-} , CO_3^{2-} , O^{2-}

a word equation

uses names: *magnesium + oxygen → magnesium oxide*

a symbol equation

uses formulae, and must be **balanced**

state symbols

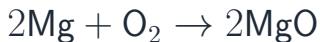
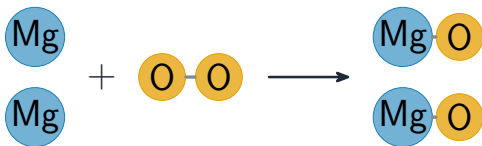
(s) solid, (l) liquid, (g) gas, (aq) **aqueous** – dissolved in water

an ionic equation

shows only the ions that actually change; spectator ions are left out

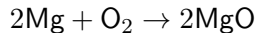
Balance by changing **the numbers in front**, never the formulae themselves. Changing a subscript makes it a different substance.

Writing and balancing equations



2 Mg and 2 O on each side: balanced

A **balanced** symbol equation has the **same** number of each atom on both sides.



State symbols : (s) (l) (g) (aq).

The words an equation uses

Reactants

The **reactants** are on the left – what you start with.

Products

The **products** are on the right – what you finish with.

Balancing

Change the *numbers in front*, never the formulas. Atoms are conserved, so each element must appear the same number of times on both sides.

In solution

A **solute** dissolves in a solvent to give a solution; concentration is moles of solute per dm^3 .

A balanced equation is a recipe in *moles*, which is why every reacting-mass calculation begins by writing one.



The mole and molar mass

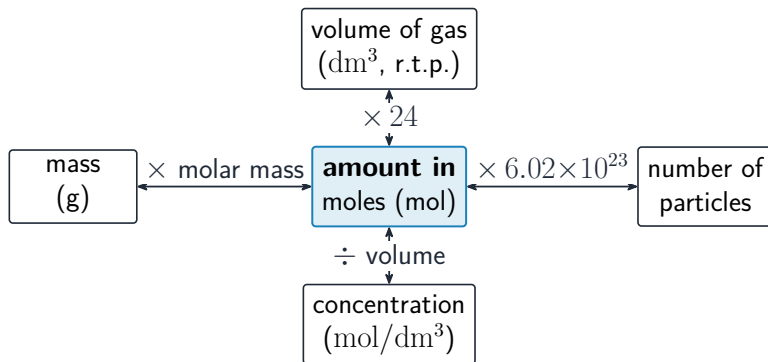
One **mole** = 6.02×10^{23} particles (the **Avogadro constant**).

$$n = \frac{\text{mass}}{\text{molar mass}}$$

Example

36 g of water: $n = \frac{36}{18} = 2 \text{ mol.}$

The mole and molar mass, in detail



One **mole** = 6.02×10^{23} particles (the **Avogadro constant**).

Gas volumes and concentration

At r.t.p. one mole of **any** gas fills 24 dm^3 .

$$V = n \times 24$$

Example

0.25 mol CO_2 : $V = 0.25 \times 24 = 6 \text{ dm}^3$.

$$\text{conc (mol/dm}^3\text{)} = \frac{n}{V(\text{dm}^3)}$$

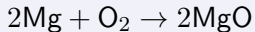
Convert volume first: $1 \text{ dm}^3 = 1000 \text{ cm}^3$.

Reaction calculations

Three steps:

- 1 known mass $\rightarrow$ moles
- 2 use the equation **ratio**
- 3 moles $\rightarrow$ mass / volume

12 g Mg burns: mass of MgO?



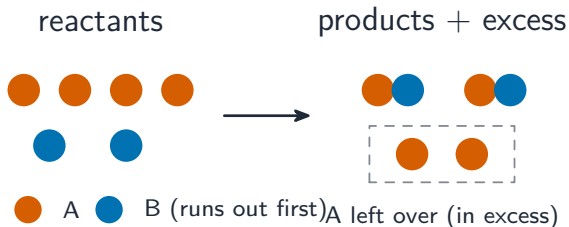
$$\text{moles Mg} = 12/24 = 0.5$$

$$\text{ratio } 1:1 \Rightarrow 0.5 \text{ mol MgO}$$

$$\text{mass} = 0.5 \times 40 = \mathbf{20 \text{ g}}$$

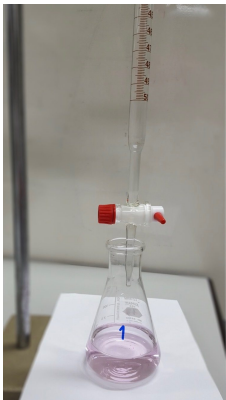


Limiting reactant



The **limiting reactant** runs out first and **decides** how much product forms. The other reactant is left over **in excess**.

Titration



**25.0 cm³ NaOH neutralised by
20.0 cm³ of 0.10 HCl**

$$\text{moles HCl} = 0.10 \times \frac{20.0}{1000} = 0.0020$$

$$\text{ratio 1:1} \Rightarrow 0.0020 \text{ mol NaOH}$$

$$\text{conc} = \frac{0.0020}{25.0/1000} = \mathbf{0.08 \text{ mol/dm}^3}$$

A titration measures exactly how much acid reacts with a base.

Empirical formula and percentages

40% Ca, 12% C, 48% O

$\div A_r$: 1, 1, 3

ratio 1:1:3 $\Rightarrow$ **CaCO₃**

$$\% \text{ by mass} = \frac{\text{mass of element}}{M_r} \times 100$$

% N in NH₄NO₃

$$\frac{28}{80} \times 100 = \mathbf{35\%}$$

Relative masses



A laboratory balance measures mass – the basis of mole calculations.

The masses, and the amount of substance

Relative atomic mass

A_r : the average mass of an atom of an element, against $\frac{1}{12}$ of a carbon-12 atom.

Relative molecular mass

Relative molecular mass M_r : the A_r values of every atom in the molecule, added.

Relative formula mass

Relative formula mass: the same sum for something with no molecules – an ionic compound such as NaCl.

Amount of substance

The **amount of substance** is measured in *moles*: $n = m/M_r$, and one mole is 6.02×10^{23} particles.

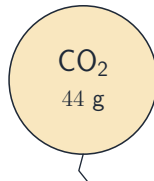
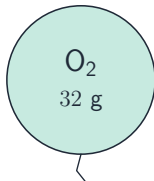
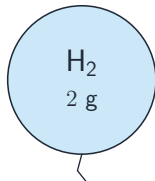
Molar gas volume

The **molar gas volume**: one mole of any gas occupies 24 dm^3 at room temperature and pressure.

Percentage composition by mass is $\frac{\text{mass of that element}}{M_r} \times 100$ – the same arithmetic, read the other way.

Volumes of gases

one mole of any gas at r.t.p.

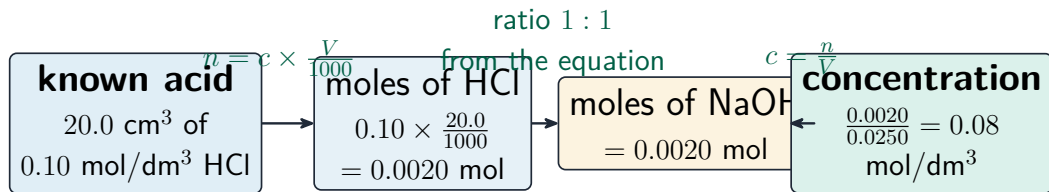


different masses, the **same volume**: 24 dm³ each

$$\text{volume (dm}^3\text{)} = \text{moles} \times 24$$

One mole of any gas fills 24 cubic decimetres at room temperature and pressure

Titration calculation



it is in 25.0 cm³ of the NaOH solution, so divide by 0.0250 dm³

The titration calculation: known solution to moles, ratio, then the unknown concentration

Exam tips

- Always turn masses, gas volumes and concentrations into **moles** first, use the balanced equation's ratio, then convert back. Never compare masses directly across an equation.
- Key formulas: $n = \frac{\text{mass}}{M_r}$, gas volume = $n \times 24 \text{ dm}^3$ at r.t.p., and concentration = $\frac{n}{\text{volume in dm}^3}$. Change cm^3 to dm^3 by dividing by 1000.
- To balance a symbol equation you may only change the big numbers *in front* of a formula, never the small subscripts inside it.
- The **limiting reactant** is the one that runs out (fewest moles once you allow for the ratio); it sets how much product forms. The other reactant is in excess.
- **Percentage yield** = $\text{actual} \div \text{theoretical} \times 100$; **percentage purity** = $\text{mass of pure substance} \div \text{mass of sample} \times 100$. Both are always 100% or less.



Topic 3 – key takeaways

1

Mole

$$n = \text{mass} / \text{molar mass}; 6.02 \times 10^{23}$$

2

Gas & conc

$$24 \text{ dm}^3 \text{ per mole; conc} = n/V$$

3

Calculations

to moles, use the ratio, back again

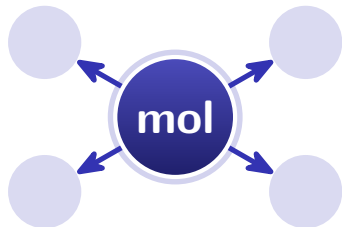
4

Limiting

the reactant that runs out sets the yield

Check yourself

- 1 How many moles are in 18 g of water?
- 2 What volume does 1 mol of gas fill at r.t.p.?
- 3 What is a spectator ion?
- 4 How do you find an empirical formula from masses?



Answers 1. 1 mol 2. 24 dm³ 3. an ion unchanged on both sides 4. divide each mass by A_r , then by the smallest