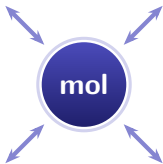


Atoms, Molecules & Stoichiometry

A-Level Chemistry ■ Topic 2

A-Level Chemistry



What we will cover

- 1 Relative masses
- 2 The mole
- 3 Formulas & equations
- 4 Reacting masses & yield
- 5 Gas volumes
- 6 Titration



mass is the basis



Relative masses

Masses are compared with $\frac{1}{12}$ of a ^{12}C atom (the unified atomic mass unit, u).

A_r

average mass
of an element's
atoms – weighted
over its isotopes

M_r

sum of the A_r
values of every atom
in the molecule

formula mass

the same idea for
an ionic compound

Chlorine is 75% ^{35}Cl and 25% ^{37}Cl

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

Four relative masses, one reference atom

Relative isotopic mass

The mass of one atom of a *single* isotope, against $\frac{1}{12}$ of a carbon-12 atom.

Relative atomic mass

A_r : the weighted average over the isotopes as they occur – which is why it is rarely a whole number.

Relative molecular mass

M_r of a **molecule**: the sum of the relative atomic masses of all its atoms.

Relative formula mass

The same idea for a substance not made of molecules – an ionic compound, where there is no molecule to weigh.

The Avogadro constant

$6.02 \times 10^{23} \text{ mol}^{-1}$: how many particles are in a mole, and what turns any of these masses into a count.

All four are ratios against the same standard, so they carry **no units**. Only the Avogadro constant does.

The mole

$$n = \frac{m}{M}$$

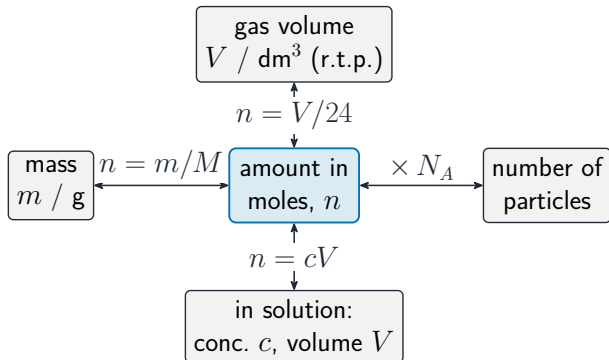
$$N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$$

Worked example

moles in 8.0 g of CH_4 ($M = 16$): $n = \frac{8.0}{16} = 0.50 \text{ mol}$

One **mole** contains N_A particles (atoms, molecules or ions); the **molar mass** M in g mol^{-1} equals A_r or M_r .

The mole map



A **compound** is two or more elements chemically joined. Chemists count its particles in **moles**, as we count eggs in dozens – the hub of every amount calculation:

- mass: $n = m/M$
- particles: $\times N_A$
- gas: $n = V/24$
- solution: $n = cV$



Formulas and equations

empirical formula

simplest ratio – CH_3

molecular formula

actual atoms – C_2H_6

Equations are **balanced** with **state symbols** (s, l, g, aq). An **ionic equation** drops **spectator ions** : $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$.

Formulas and equations, in detail

element	% by mass	$\div A_r \rightarrow$ moles	smallest (3.33)	ratio
C	40.0	$40.0/12 = 3.33$	1.0	1
H	6.7	$6.7/1 = 6.7$	2.0	2
O	53.3	$53.3/16 = 3.33$	1.0	1

empirical formula CH_2O - if $M_r = 180$, it fits $180/30 = 6$ times: molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$

The empirical-formula recipe: divide by A_r , divide by the smallest, read off the ratio

Hydrated and anhydrous solids

Some solids hold water inside their crystals.

This water is the **water of crystallisation** .

Reacting masses and yield

Worked example: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

$4.8 \text{ g Mg} = 0.20 \text{ mol}$; ratio $1:1 \Rightarrow 0.20 \text{ mol MgO} = 0.20 \times 40 = 8.0 \text{ g}$

$$\text{percentage yield} = \frac{\text{actual product}}{\text{maximum possible}} \times 100\%$$

The numbers in a balanced equation give the mole ratio – the **stoichiometry** .

Limiting reagent



start with



4 H₂ 1 O₂

O₂ runs out first:
limiting reagent

you get



2 H₂O made

2 H₂ left over (excess)

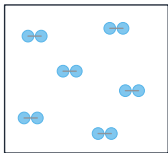
The **limiting reagent** runs out first and decides how much product forms.

The leftover reactant is in **excess**. Always base the calculation on the limiting reagent.

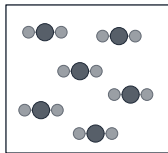


Gas volumes

same volume, same temperature and pressure



hydrogen, H_2



carbon dioxide, CO_2

the same **number** of molecules (6 each), even though CO_2 is bigger

Equal volumes of any gases
(same temperature & pressure)
hold equal numbers of molecules.

At r.t.p., one mole of any
gas is 24.0 dm^3 , so $n =$
 $\frac{V}{24.0}$.

The words a solution question uses

Solute and solution

Concentration is the amount of **solute** in each cubic decimetre of **solution**, in mol dm^{-3} : $c = n/V$.

State symbols

(s) solid, (l) liquid, (g) gas, and (aq) **aqueous** – dissolved in water. An equation without them is incomplete.

Excess reagent

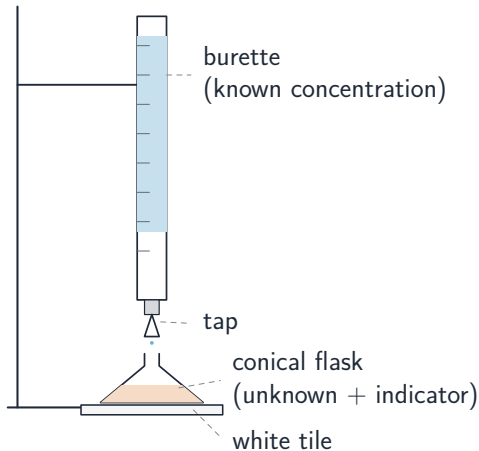
The **limiting reagent** runs out first and fixes the yield; whatever is left over is the **excess reagent**.

Hydrocarbons

For burning **hydrocarbons** – compounds of carbon and hydrogen only – compare *gas volumes* directly, since equal volumes hold equal moles.

Convert to moles first, every time. Mass, concentration and gas volume are three doors into the same room.

Titration



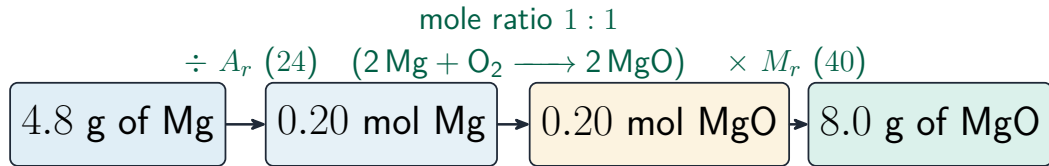
Concentration : $n = c \times V$ (in mol dm^{-3} , V in dm^3).

NaOH neutralised by HCl

20.0 cm^3 of 0.100 mol dm^{-3} HCl $\Rightarrow$
NaOH = 0.0800 mol dm^{-3}

A **titration** finds an unknown concentration.

Reacting masses and percentage yield



mass $\rightarrow$ moles $\rightarrow$ use the equation's ratio $\rightarrow$ moles $\rightarrow$ mass

The mole bridge: mass to moles, ratio, then back to mass

Relative masses of atoms and molecules



$$A_r = \frac{(75 \times 35) + (25 \times 37)}{100} = 35.5$$

Relative atomic mass is a **weighted** average over the isotopes – chlorine is 75% ^{35}Cl and 25% ^{37}Cl , so $A_r = 35.5$. **Atoms** are far too light to weigh in grams, so every mass is compared to $\frac{1}{12}$ of a ^{12}C atom – the **unified atomic mass unit**, u.

Predicting an ion's charge from the group

In an **ionic compound** the positive and negative charges balance exactly, so the compound is neutral overall.

Group	1	2	13	15	16	17
Ion charge	+1	+2	+3	-3	-2	-1

Hydrogen forms H^+ ; the Group 18 noble gases do not normally form ions at all. Where a metal has more than one charge, a Roman numeral gives the **oxidation number** : iron(II) is Fe^{2+} , iron(III) is Fe^{3+} .

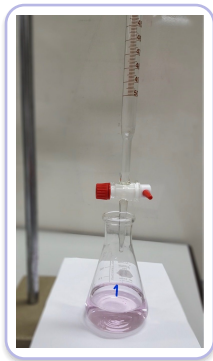
Worked example – the A_r of copper

A mass spectrometer shows copper is 69.2% ^{63}Cu and 30.8% ^{65}Cu

$$A_r = \frac{(69.2 \times 63) + (30.8 \times 65)}{100} = \frac{4359.6 + 2002.0}{100} = 63.6$$

Multiply each isotope's mass by its abundance, add, divide by 100. The answer always sits **between** the isotope masses, nearer the commoner one – which is how you check it.

Reacting masses and volumes



A titration finds reacting volumes precisely.

Exam tips

- Convert volumes to dm^3 **before** using concentration ($25.0 \text{ cm}^3 = 0.0250 \text{ dm}^3$); the missing $\div 1000$ is the most common titration error.
- Use the **balancing numbers** as the mole ratio between species – never the M_r values.
- Empirical formula: divide each mass/percentage by A_r , then by the **smallest**, then scale to whole numbers; use M_r to reach the molecular formula.
- For gases at r.t.p. use $\text{volume} = \text{moles} \times 24 \text{ dm}^3$; quote the equation you use every time.
- Give the answer to the same **significant figures** as the data (usually 3) and always include units.



Topic 2 – key takeaways

1

The mole

$$n = m/M; N_A = 6.02 \times 10^{23}$$

2

Formulas

empirical (ratio) vs molecular (actual)

3

Reacting masses

use the mole ratio; limiting reagent

4

Gases & solutions

$$n = V/24 \text{ at r.t.p.}; n = cV$$

Check yourself

- 1 How many moles in 16 g of CH_4 ?
- 2 Empirical formula of C_2H_6 ?
- 3 What volume does 1 mol of gas occupy at r.t.p.?
- 4 Which reagent decides how much product forms?



Answers 1. 1 mol 2. CH_3 3. 24.0 dm^3 4. the limiting reagent