

Atoms, Molecules & Stoichiometry

A-Level Chemistry ■ Topic 2

A-Level Chemistry



Atoms, Molecules & Stoichiometry

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

1

Masses and the mole

Atoms, Molecules & Stoichiometry

└─ Masses and the mole

Relative masses

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

A_r

相对原子质量
average mass
of an element's
atoms – weighted
over its isotopes

M_r

相对分子质量
sum of the A_r
values of every atom
in the molecule

formula mass

相对式量

the same idea for
an ionic compound

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

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

Atoms, Molecules & Stoichiometry

└─ Masses and the mole

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** 分子: 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.

Atoms, Molecules & Stoichiometry

└─ Masses and the mole

The mole

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

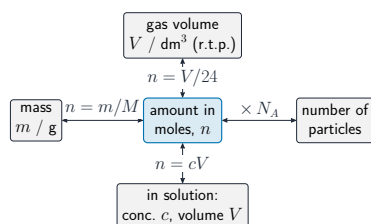
$$N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$$

Worked example

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

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

The mole map



A **compound** 化合物 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: $n = cV$

2

Formulas and reacting masses

Formulas and equations

empirical formula 实验式

simplest ratio – CH_3

molecular formula 分子式

actual atoms – C_2H_6

Equations are **balanced** 配平 with **state symbols** 状态符号 (s, l, g, aq). An **ionic equation** 离子方程式 drops **spectator ions** 旁观离子: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$.

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** 结晶水.

Reacting masses and yield

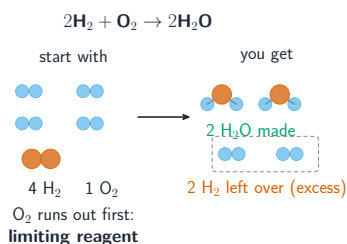
Worked example: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

4.8 g Mg = 0.20 mol; ratio 1:1 $\Rightarrow$ 0.20 mol MgO = $0.20 \times 40 = 8.0$ 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** 化学计量.

Limiting reagent

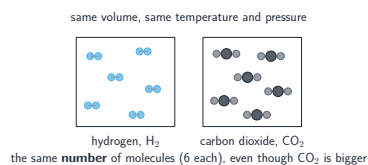


The **limiting reagent** 限量试剂 runs out first and decides how much product forms.

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

3 Gases and solutions

Gas volumes



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 dm³, so $n = \frac{V}{24.0}$.

The words a solution question uses

Solute and solution Concentration is the amount of **solute** 溶质 in each cubic decimetre of **solution** 溶液, in mol dm⁻³: $c = n/V$.

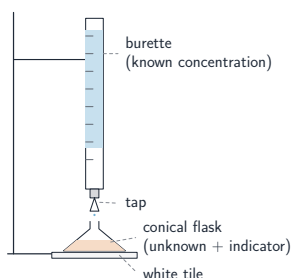
State symbols (s) solid, (l) liquid, (g) gas, and (aq) **aqueous** 水溶液 – dissolved in water. An equation without them is incomplete.

Excess reagent The **limiting reagent** 限量试剂 runs out first and fixes the yield; whatever is left over is the **excess reagent** 过量试剂.

Hydrocarbons For burning **hydrocarbons** 碳氢化合物 – 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



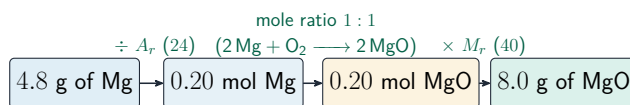
Concentration 浓度: $n = c \times V$ (in mol dm⁻³, V in dm³).

NaOH neutralised by HCl

20.0 cm³ of 0.100 mol dm⁻³ HCl ⇒
NaOH = 0.0800 mol dm⁻³

A **titration** 滴定 finds an unknown concentration.

Reacting masses and percentage yield



mass → moles → use the equation's ratio → moles → 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** 原子 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** 统一原子质量单位, u.

Predicting an ion's charge from the group

In an **ionic compound** 离子化合物 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** 氧化数: iron(II) is Fe^{2+} , iron(III) is Fe^{3+} .

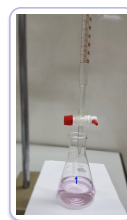
Worked example – the A_r of copper

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 volume = 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.

4

Recap

Topic 2 – key takeaways

1 The mole

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

3 Reacting masses

use the mole ratio; limiting reagent

2 Formulas

empirical (ratio) vs molecular (actual)

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