Stoichiometry

IGCSE Chemistry

Stoichiometry 化学计量 is the study of the amounts of substances in a reaction — how much reacts and how much is made. This topic is mostly about counting atoms and doing calculations.

Chemical formulae

A **formula** 化学式 shows which **atoms** 原子 are in a substance, and how many of each. There are two kinds of formula you must know.

- The **molecular formula** 分子式 is the actual number of each atom in one **molecule** 分子. For glucose it is $C_6H_{12}O_6$.
- The **empirical formula** 实验式 is the simplest whole-number **ratio** 比例 of the atoms or **ions** 离子 in a **compound** 化合物. For glucose it is CH_2O.

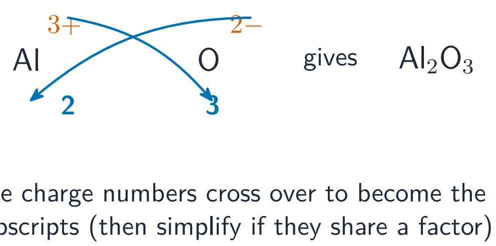
Working out a formula

If you are given a model or diagram, just count the atoms of each element and write them as a formula.

For an **ionic compound** 离子化合物 you can work out the formula from the charges on the ions. The total positive charge must balance the total negative charge, because the compound has no overall charge. Some common ions:

Positive ions	**Negative ions**
Na^+, K^+, H^+	Cl^-, OH^-, NO_3^-
Mg^{2+}, Ca^{2+}, Cu^{2+}	O^{2-}, SO_4^{2-}, CO_3^{2-}
Al^{3+}	N^{3-}

For example, Na^+ and O^{2-}: you need two Na^+ to balance one O^{2-}, so the formula is Na_2O.



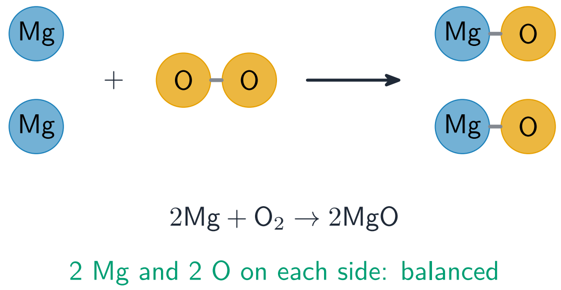
For an ionic compound, the ion charges cross over to give the formula (here Al_2O_3)

Writing equations

An equation shows how **reactants** 反应物 (the starting substances) change into **products** 生成物 (the substances made).

- A **word equation** 文字方程式 uses names: *magnesium + oxygen → magnesium oxide*.
- A **symbol equation** 化学方程式 uses formulae and must be **balanced** 配平—the same number of each atom on both sides.

$$2\mathrm{Mg} + \mathrm{O_2} \rightarrow 2\mathrm{MgO}$$



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

You add **state symbols** 状态符号 in brackets to show the physical state: (s) solid, (l) liquid, (g) gas, and (aq) for **aqueous** 水溶液 (dissolved in water).

$$\mathrm{Zn}(s) + 2\mathrm{HCl}(aq) \rightarrow \mathrm{ZnCl_2}(aq) + \mathrm{H_2}(g)$$

Ionic equations

An **ionic equation** 离子方程式 shows only the ions that actually change. Ions that are the same on both sides are **spectator ions** 旁观离子 and are left out. For example, when an acid reacts with an alkali:

$$\mathrm{H^+}(aq) + \mathrm{OH^-}(aq) \rightarrow \mathrm{H_2O}(l)$$

Relative masses

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

The **relative atomic mass** 相对原子质量 (A_r) of an **element** 元素 is the average **mass** 质量 of its atoms compared to $\frac{1}{12}$ of the mass of one ^{12}C atom.

The **relative molecular mass** 相对分子质量 (M_r) is the sum of the relative atomic masses of all the atoms in the molecule. For ionic compounds we use the **relative formula mass** 相对式量, found the same way from the formula.

$$M_r(\mathrm{H_2O}) = (2 \times 1) + 16 = 18$$

Reacting masses by simple proportion

You can sometimes find a reacting mass without the mole. If 24 g of magnesium makes 40 g of magnesium oxide, then 12 g of magnesium (half as much) makes 20 g of magnesium oxide.

The mole

The **mole** 摩尔 (symbol mol) is the unit for the **amount of substance** 物质的量. One mole of any substance contains 6.02×10^{23} **particles** 粒子 (atoms, ions or molecules). This number is the **Avogadro constant** 阿伏伽德罗常数.

The **molar mass** 摩尔质量 is the mass of one mole, in grams per mole (g/mol). Its number is the same as the A_r or M_r. The key relationship is:

$$\text{amount (mol)} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

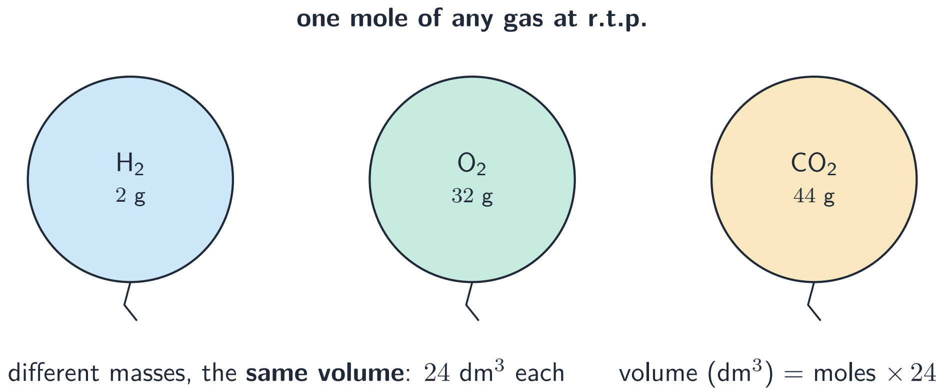
Worked example. How many moles are in 36 g of water? Molar mass of water = 18 g/mol.

$$n = \frac{36}{18} = 2 \text{ mol}$$

To find the number of particles, multiply the moles by the Avogadro constant.

Volumes of gases

At room temperature and pressure (r.t.p.), one mole of any gas takes up the same **volume** 体积. This **molar gas volume** 摩尔气体体积 is 24 dm^3.



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

$$\text{volume of gas (dm}^3) = \text{amount (mol)} \times 24$$

Worked example. What volume does 0.25 mol of carbon dioxide occupy at r.t.p.?

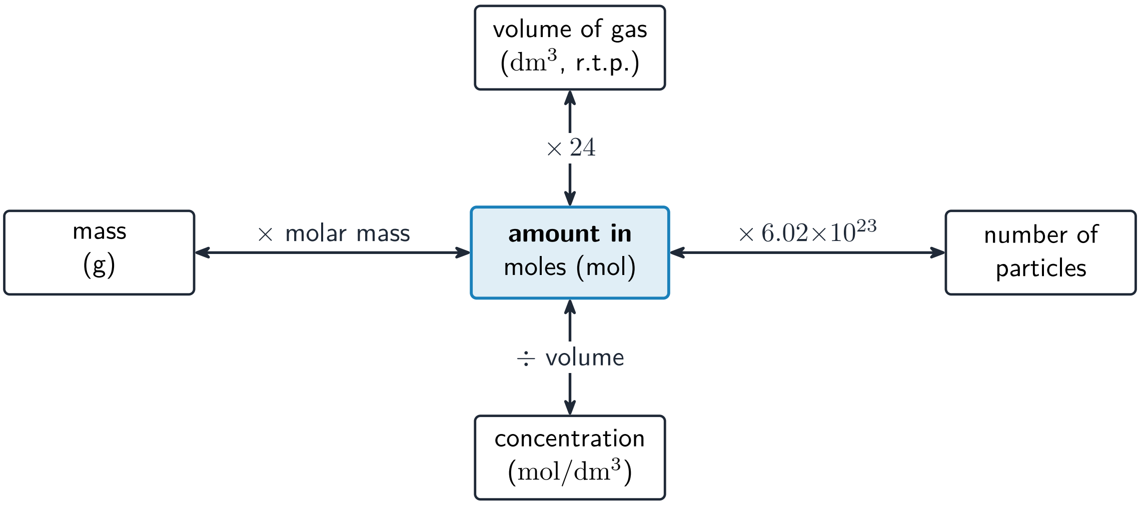
$$\text{volume} = 0.25 \times 24 = 6 \text{ dm}^3$$

Concentration of solutions

The **concentration** 浓度 of a **solution** 溶液 can be given in g/dm^3 or in mol/dm^3.

$$\text{concentration (mol/dm}^3) = \frac{\text{amount of solute (mol)}}{\text{volume (dm}^3)}$$

Remember to convert volume: 1 dm^3 = 1000 cm^3, so divide a volume in cm^3 by 1000.



Moles sit at the centre: multiply going outwards, divide coming back (concentration is moles ÷ volume)

Doing reaction calculations

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

Most calculations follow the same steps: change the known amount into moles, use the balanced equation to find the moles of what you want, then change back into mass, volume or concentration.

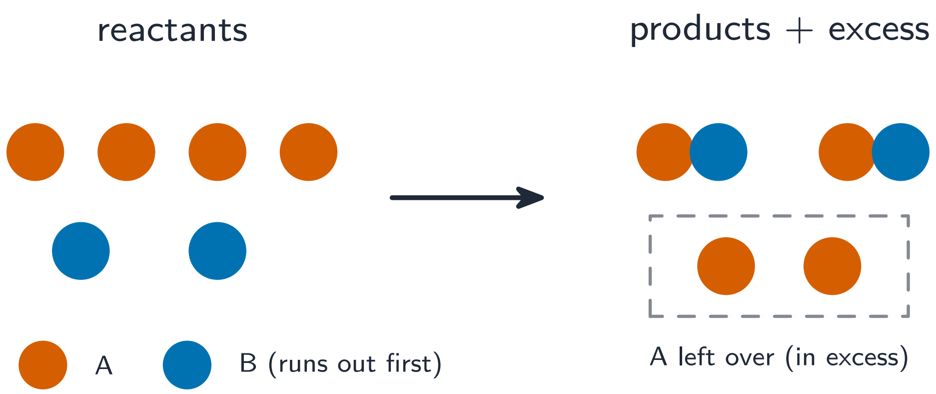
Worked example. What mass of magnesium oxide forms when 12 g of magnesium burns completely? $2Mg + O_2 \rightarrow 2MgO$. ($A_r$: Mg = 24, O = 16.)

Moles of Mg = 12/24 = 0.5 mol. The ratio Mg : MgO is 1 : 1, so 0.5 mol of MgO forms.

Its molar mass is $24 + 16 = 40$ g/mol, so the mass is $0.5 \times 40 = 20$ g —the same answer as the simple-proportion method above.

Limiting reactant

When two reactants are mixed, often one runs out first. This is the **limiting reactant** 限量反应物. It decides how much product can form; any other reactant is in excess (left over).



B runs out first, so it is the limiting reactant and sets how much product forms; the spare A is left over in excess

Worked example. 8 g of hydrogen reacts with 8 g of oxygen: $2H_2 + O_2 \rightarrow 2H_2O$. Which is the limiting reactant, and what mass of water forms? (A_r: H = 1, O = 16.)

Moles of $H_2 = 8/2 = 4$ mol; moles of $O_2 = 8/32 = 0.25$ mol. The equation needs $2H_2$ for every one O_2, so 0.25 mol of O_2 reacts with only 0.5 mol of H_2. There is far more hydrogen than that, so **oxygen is the limiting reactant** and the hydrogen is in excess. The oxygen makes $2 \times 0.25 = 0.5$ mol of water, so the mass is $0.5 \times 18 = 9$ g. Always work in moles, not masses —the reactant with the smaller mass is not always the one that runs out.

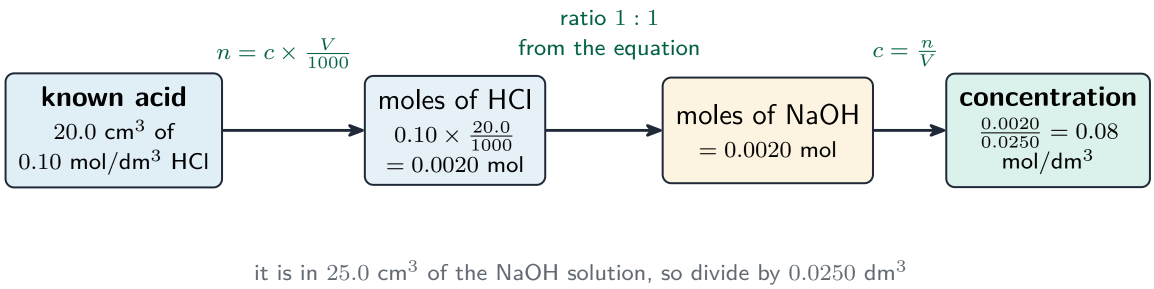
Titration calculation

In a **titration** 滴定 you measure the volume of one solution that reacts with another. From the volume and concentration you find the moles of one **solute** 溶质, then use the equation ratio to find the moles, concentration or volume of the other.

Worked example. 25.0 cm^3 of sodium hydroxide solution is exactly neutralised by 20.0 cm^3 of 0.10 mol/dm^3 hydrochloric acid. $NaOH + HCl \rightarrow NaCl + H_2O$. Find the concentration of the sodium hydroxide.

Moles of HCl $= 0.10 \times \dfrac{20.0}{1000} = 0.0020$ mol. The ratio is $1:1$, so there are 0.0020 mol of NaOH in 25.0 cm^3:

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



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

Empirical and molecular formulae from data

To find an empirical formula from masses or percentages:

1. Divide each element's mass (or %) by its A_r.
2. Divide all the answers by the smallest one to get the simplest ratio.

To find the molecular formula, compare the empirical formula mass with the real M_r and multiply up.

Worked example. A compound contains 40% calcium, 12% carbon and 48% oxygen by mass. Find its empirical formula. (A_r: Ca = 40, C = 12, O = 16.)

Divide each percentage by its A_r: Ca $= 40/40 = 1$, C $= 12/12 = 1$, O $= 48/16 = 3$. The ratio is 1 : 1 : 3, so the empirical formula is $CaCO_3$.

Percentages

- **Percentage yield 产率** compares how much product you actually got with the most you could get: $\dfrac{\text{actual}}{\text{theoretical}} \times 100$.
- **Percentage composition by mass 质量分数** of an element $= \dfrac{\text{mass of the element}}{M_r} \times 100$.
- **Percentage purity 纯度** $= \dfrac{\text{mass of pure substance}}{\text{mass of impure sample}} \times 100$.

Worked example. Find the percentage by mass of nitrogen in ammonium nitrate, NH_4NO_3. (A_r: N = 14, H = 1, O = 16.)

$M_r = (2 \times 14) + (4 \times 1) + (3 \times 16) = 80$. The mass of nitrogen is $2 \times 14 = 28$, so

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

Worked example. A reaction could make at most 5.0 g of product, but only 4.0 g is actually collected. Find the percentage yield.

$$\% \text{ yield} = \frac{4.0}{5.0} \times 100 = 80\%$$

Some product is always lost —left behind in the apparatus, or used up in side reactions —so the yield is almost never 100%.

Worked example. A 50 g sample of impure calcium carbonate contains 45 g of pure calcium carbonate. Find its percentage purity.

$$\% \text{ purity} = \frac{45}{50} \times 100 = 90\%$$

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 $= n \times 24$ 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** = actual ÷ theoretical × 100; **percentage purity** = mass of pure substance ÷ mass of sample × 100. Both are always 100% or less.