# The Mole Concept

Matthew Williams • Chemistry • May 15, 2026

## The Mole Concept

Atoms and molecules are impossibly small — far too tiny to count or weigh individually. The mole is the bridge between the microscopic world of particles and the macroscopic quantities measured in the lab.

### The Mole and Avogadro's Constant

A **mole** is the amount of substance that contains $6.02 \times 10^{23}$ particles. That number — $6.02 \times 10^{23}$ — is **Avogadro's constant** ( $N_A$ ). The particles can be atoms, molecules, ions, or electrons, depending on the substance.

One mole of any substance contains the same number of particles, regardless of what the substance is. One mole of carbon atoms, one mole of water molecules, and one mole of sodium ions all contain $6.02 \times 10^{23}$ particles.



### Molar Mass

**Molar mass** is the mass of one mole of a substance, expressed in g mol$^{-1}$. Numerically, the molar mass of an element equals its relative atomic mass in grams. For a compound, the molar mass equals the sum of the relative atomic masses of all atoms in one formula unit.

| Substance | Formula | Molar mass |
|---|---|---|
| Hydrogen | $H_2$ | 2 g mol$^{-1}$ |
| Carbon | C | 12 g mol$^{-1}$ |
| Water | $H_2O$ | 18 g mol$^{-1}$ |
| Sodium chloride | NaCl | 58.5 g mol$^{-1}$ |
| Calcium carbonate | $CaCO_3$ | 100 g mol$^{-1}$ |
| Sulfuric acid | $H_2SO_4$ | 98 g mol$^{-1}$ |

## Calculations Involving the Mole

The fundamental relationship between moles, mass, and molar mass is:

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

where $n$ is the number of moles, $m$ is the mass in grams, and $M$ is the molar mass in g $mol^{-1}$.

> 📋 **Example**
> How many moles are in 44 g of carbon dioxide ($CO_2$)?
> Molar mass of $CO_2$ = 12 + (2 × 16) = 44 g $mol^{-1}$
> $$n = \frac{44}{44} = 1 \text{ mol}$$

> 📋 **Example**
> What is the mass of 0.25 mol of $CaCO_3$?
> Molar mass of $CaCO_3$ = 40 + 12 + (3 × 16) = 100 g $mol^{-1}$
> $$m = n \times M = 0.25 \times 100 = 25 \text{ g}$$

## Relative Formula Mass

The **relative formula mass** ($M_r$) of a compound is the sum of the relative atomic masses of all the atoms in one formula unit. For ionic compounds the term relative formula mass is used; for molecular compounds it is often called relative molecular mass. Neither is expressed in units — they are pure numbers relative to carbon-12.

> 📋 **Example**
> Find the relative formula mass of aluminium sulfate, $Al_2(SO_4)_3$
> Al: 2 × 27 = 54
> S: 3 × 32 = 96
> O: 12 × 16 = 192
> $M_r$ = 54 + 96 + 192 = **342**

## Percentage Composition by Mass

The percentage by mass of each element in a compound is:

$$\% \text{ by mass} = \frac{\text{mass of element in one mole}}{\text{molar mass of compound}} \times 100$$

📋 **Example**

Find the percentage by mass of nitrogen in ammonium nitrate, $NH_4NO_3$

Molar mass = 14 + 4 + 14 + 48 = 80 g $mol^{-1}$
Mass of N = 14 + 14 = 28 g

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



## Empirical and Molecular Formulae

The **empirical formula** gives the simplest whole-number ratio of atoms in a compound. The **molecular formula** gives the actual number of atoms in one molecule. The molecular formula is always a whole-number multiple of the empirical formula.

To find the empirical formula from percentage composition:

- 1. Treat the percentages as masses in grams.
- 2. Divide each mass by its relative atomic mass to find moles.
- 3. Divide all mole values by the smallest to get the simplest ratio.
- 4. Round to the nearest whole number (or multiply through if fractions appear).

📋 **Example**

A compound contains 40% carbon, 6.7% hydrogen, and 53.3% oxygen by mass. Find its empirical formula.

| Element | % | ÷ Ar | Ratio |
|---|---|---|---|
| C | 40 | 40/12 = 3.33 | 3.33/3.33 = **1** |
| H | 6.7 | 6.7/1 = 6.7 | 6.7/3.33 = **2** |
| O | 53.3 | 53.3/16 = 3.33 | 3.33/3.33 = **1** |

Empirical formula: **$CH_2O$**

To find the molecular formula, divide the given molar mass by the empirical formula mass and multiply the subscripts.

## Avogadro's Law and Gas Volumes

**Avogadro's Law** states that equal volumes of all gases, measured at the same temperature and pressure, contain equal numbers of molecules.

This leads directly to the concept of molar volume. At the conditions used in CSEC:

| Conditions | Molar volume |
|---|---|
| RTP (room temperature and pressure, 25 °C and 1 atm) | 24 $dm^3$ $mol^{-1}$ |
| STP (standard temperature and pressure, 0 °C and 1 atm) | 22.4 $dm^3$ $mol^{-1}$ |

$$n = \frac{V}{24} \text{ (at RTP)}$$

📋 **Example**
What volume does 0.5 mol of oxygen gas occupy at RTP?

$$V = n \times 24 = 0.5 \times 24 = 12 \text{ dm}^3$$

## Law of Conservation of Matter

The **Law of Conservation of Matter** states that matter is neither created nor destroyed in a chemical reaction. The total mass of reactants always equals the total mass of products. This is why chemical equations must be balanced — the same atoms must appear on both sides.

## Balanced Equations and State Symbols

A balanced molecular equation shows the formulae of all reactants and products with state symbols, and has equal numbers of each type of atom on both sides.

State symbols:

- (s) — solid
- (l) — liquid

- (g) — gas
- (aq) — aqueous (dissolved in water)

**Ionic equations** show only the species that actually change in a reaction — spectator ions (ions present but unchanged) are omitted.

> 📋 **Example**
> Molecular equation:
> $\mathrm{HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)}$
>
> Full ionic equation:
> $\mathrm{H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)}$
>
> Net ionic equation (spectators $Na^+$ and $Cl^-$ removed):
> $\mathrm{H^+(aq) + OH^-(aq) \rightarrow H_2O(l)}$

## Applying the Mole Concept to Equations

The coefficients in a balanced equation are mole ratios — they tell you how many moles of each substance react or are produced. This allows calculation of reacting masses and volumes from any given starting quantity.

**Method:**

- 1. Write the balanced equation.
- 2. Identify the mole ratio between the substances of interest.
- 3. Calculate moles of the given substance.
- 4. Use the ratio to find moles of the unknown substance.
- 5. Convert to mass (using $m = nM$) or volume (using $V = n \times 24$) as required.

> 📋 **Example**
> What mass of iron(III) oxide is produced when 5.6 g of iron reacts completely with oxygen?
>
> $$\mathrm{4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)}$$
>
> Moles of Fe = 5.6 / 56 = 0.1 mol
> Ratio Fe : $Fe_2O_3$ = 4 : 2 = 2 : 1
> Moles of $Fe_2O_3$ = 0.1 / 2 = 0.05 mol
> Molar mass of $Fe_2O_3$ = (2 × 56) + (3 × 16) = 160 g mol$^{-1}$
> Mass of $Fe_2O_3$ = 0.05 × 160 = **8.0 g**

## Concentration and Standard Solutions

**Molar concentration** is the number of moles of solute dissolved per cubic decimetre (litre) of solution:

$$c = \frac{n}{V}$$

where $c$ is in mol dm⁻³ and $V$ is in $dm^3$. Note: 1 $dm^3$ = 1000 $cm^3$, so always convert $cm^3$ by dividing by 1000.

**Mass concentration** is the mass of solute dissolved per unit volume, expressed in g $dm^{-3}$.

A **standard solution** is a solution whose concentration is accurately known. Standard solutions are prepared by dissolving an accurately weighed mass of solute in a known volume of solvent using a volumetric flask.

> 📝 **Exam Tip**
> When using $n = cV$, volume must be in $dm^3$. A 25 $cm^3$ pipette reading becomes 0.025 $dm^3$. Forgetting this conversion is the single most common error in calculation questions.

## Volumetric Analysis Calculations

In a titration, one solution of known concentration (the standard solution) reacts completely with another. The unknown concentration is calculated from the volume used and the mole ratio from the balanced equation.

**General approach:**

- 1. Calculate moles of the known solution: $n = cV$
- 2. Use the mole ratio from the balanced equation to find moles of the unknown.
- 3. Calculate the concentration of the unknown: $c = n/V$

📋 **Example**

25.0 cm$^3$ of NaOH solution was neutralised by 20.0 cm$^3$ of 0.10 mol dm$^{-3}$ HCl.
Find the concentration of the NaOH.

$$\mathrm{HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)}$$

Moles HCl = 0.10 × 0.020 = 0.0020 mol
Ratio 1 : 1, so moles NaOH = 0.0020 mol
Concentration NaOH = 0.0020 / 0.025 = **0.080 mol dm$^{-3}$**

📋 **Example**

15.0 cm$^3$ of 0.60 mol dm$^{-3}$ $HNO_3$ neutralised 25.0 cm$^3$ of $Na_2CO_3$ solution.

$$\mathrm{2HNO_3 + Na_2CO_3 \rightarrow 2NaNO_3 + H_2O + CO_2}$$

Moles $HNO_3$ = 0.60 × 0.015 = 0.0090 mol
Ratio $HNO_3$ : $Na_2CO_3$ = 2 : 1
Moles $Na_2CO_3$ = 0.0090 / 2 = 0.0045 mol
Concentration $Na_2CO_3$ = 0.0045 / 0.025 = **0.18 mol dm$^{-3}$**