

Chemical Reactions

AP Chemistry

Recognizing a Chemical Reaction

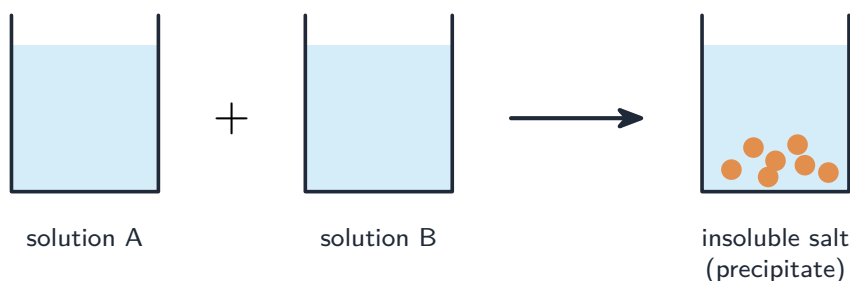


Magnesium burning: a chemical reaction rearranges atoms and releases energy as light and heat

A **chemical reaction** rearranges atoms into new substances. Signs one has happened: a color change, a gas or **precipitate** forming, or a temperature change. Atoms are conserved, so an equation must be **balanced** – the same count of each element on both sides.

Net Ionic Equations

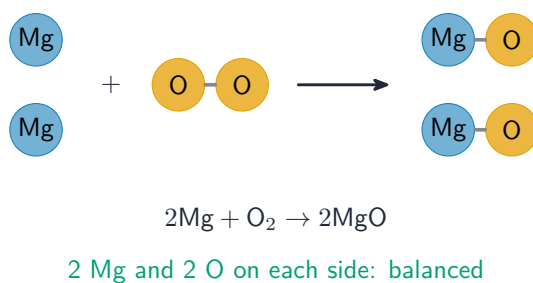
For reactions in water, ionic compounds split into ions. A **net ionic equation** shows only the species that actually change, leaving out the **spectator ions** that appear unchanged on both sides. It captures the real chemistry (e.g. $\text{Ag}^+ + \text{Cl}^- \rightarrow \text{AgCl}(s)$).



Mixing two solutions can form an insoluble precipitate

Three Ways to Represent a Reaction

The same reaction can be shown as a **symbolic** equation, a **particulate** drawing (atoms and molecules), and a **macroscopic** observation (what you see). Moving between these levels – connecting the equation to the particles to the beaker – is a core skill.



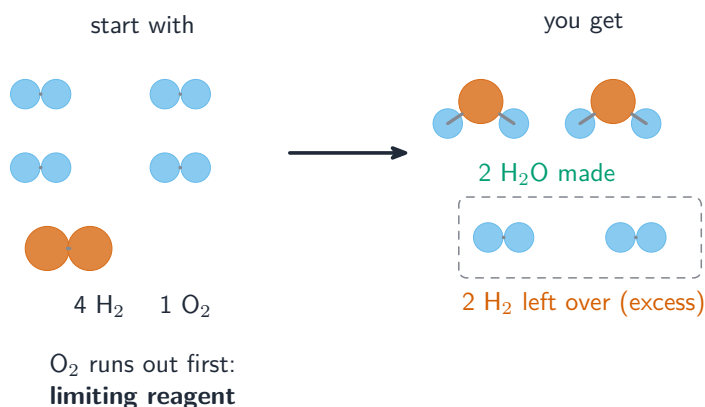
A balanced equation has the same number of each atom on both sides

Physical Changes versus Chemical Changes

A **physical change** alters form but not identity (melting, dissolving); a **chemical change** makes new substances by breaking and forming bonds. Dissolving salt is physical; the salt is unchanged and recoverable.

Stoichiometry

Stoichiometry uses the balanced equation's mole ratios to relate amounts of reactants and products. The path is always grams → moles → (mole ratio) → moles → grams. The **limiting reactant** runs out first and sets the maximum product (the **theoretical yield**); the **percent yield** compares actual to theoretical.

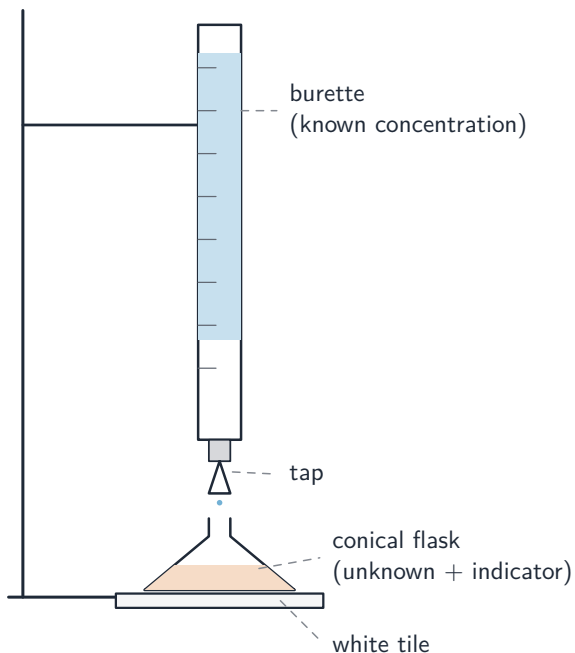


Worked example. How much water forms when 4.0 g of hydrogen burns in excess oxygen? $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$. Convert to moles, cross the mole ratio ($2 : 2 = 1 : 1$ here), convert back:

$$n(\text{H}_2) = \frac{4.0}{2.0} = 2.0 \text{ mol} \Rightarrow n(\text{H}_2\text{O}) = 2.0 \text{ mol} \Rightarrow m = 2.0 \times 18 = 36 \text{ g}.$$

Worked example (limiting reactant). 10.0 g of N_2 reacts with 5.0 g of H_2 ($\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$). Which runs out? Moles: $n(\text{N}_2) = 10.0/28 = 0.36$, $n(\text{H}_2) = 5.0/2 = 2.5$. The reaction needs 3 H_2 per N_2 ; you have $2.5/0.36 = 7.0$, far more than 3, so **N_2 is limiting**. It makes $2 \times 0.36 = 0.72$ mol of ammonia, with hydrogen left over.

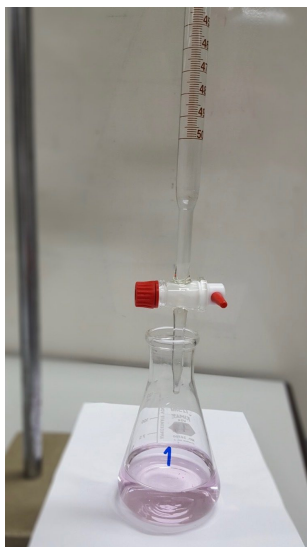
A **titration** finds an unknown concentration by reacting it with a solution of known concentration until they reach the **equivalence point** (stoichiometrically equal). From the known volume and concentration, use the mole ratio to find the unknown. An **indicator** or a pH curve signals the endpoint.



Titration apparatus: a burette delivers a known solution into a conical flask

Worked example. 25.0 mL of hydrochloric acid is exactly neutralized by 30.0 mL of 0.100 M NaOH. Find the acid's concentration. The reaction is 1 : 1, so the moles match:

$$n(\text{NaOH}) = 0.0300 \times 0.100 = 3.00 \times 10^{-3} \text{ mol} = n(\text{HCl}), \quad [\text{HCl}] = \frac{3.00 \times 10^{-3}}{0.0250} = 0.120 \text{ M}.$$



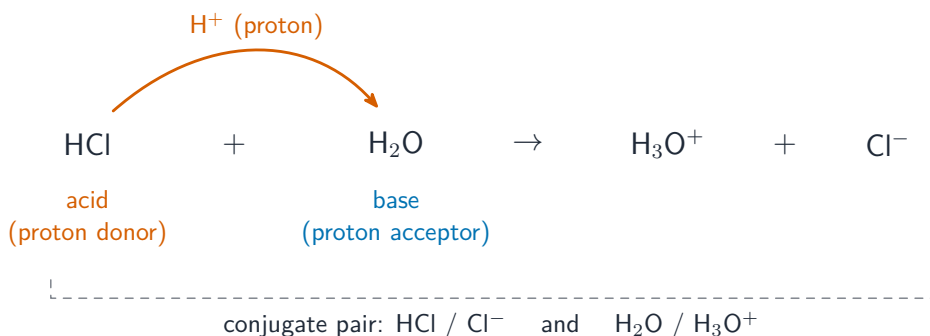
A titration setup: reacting volumes from a burette find the concentration of an unknown solution

Types of Chemical Reactions

Common patterns include **synthesis** (combining), **decomposition** (breaking apart), **combustion** (with oxygen, releasing energy), **precipitation** (forming an insoluble solid), **acid–base** (proton transfer), and **redox** (electron transfer). Recognizing the type helps predict the products.

Acid-Base Reactions

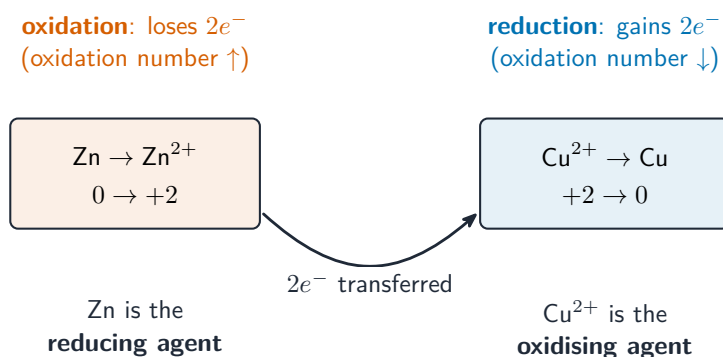
An **acid** donates a proton (H^+); a **base** accepts one (Brønsted–Lowry). They react to form water and a salt: $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$. Strong acids and bases dissociate completely; weak ones only partly.



Brønsted–Lowry: an acid donates a proton to a base, forming two conjugate pairs

Oxidation-Reduction (Redox) Reactions

In a **redox** reaction, electrons transfer between species. **Oxidation** is loss of electrons (oxidation number rises); **reduction** is gain (oxidation number falls). Track changes with **oxidation numbers**, and remember every oxidation is paired with a reduction – the electrons lost equal the electrons gained.



Redox is electron transfer: the reducing agent is oxidised, the oxidising agent reduced

Worked example. Find the oxidation number of manganese in permanganate, KMnO_4 . Potassium is +1 and each oxygen is -2 (four of them, -8). The whole formula is neutral, so $(+1) + \text{Mn} + (-8) = 0$, giving $\text{Mn} = +7$ – its maximum, which is why permanganate is a powerful oxidizing agent (it can only gain electrons).

Exam tips

- Balance every equation and work in **moles** – convert grams $\rightarrow$ moles, cross the mole ratio, then convert back.
- Find the **limiting reactant** by comparing mole ratios; it sets the maximum product (theoretical yield).
- In redox, remember **OIL RIG**: oxidation is loss of electrons, reduction is gain; every oxidation is paired with a reduction.
- For titrations use the balanced ratio to link the known and unknown at the **equivalence point**.
- A **net ionic equation** shows only the species that change; leave out spectator ions.