4 · Chemical changes

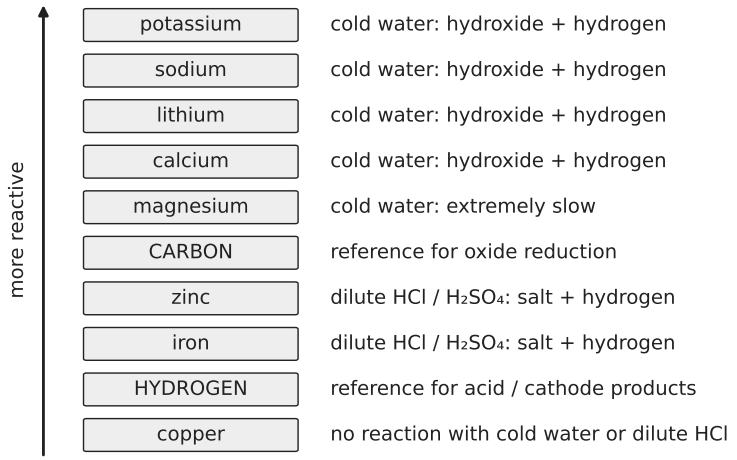
AQA · GCSE · Chemistry

Chemical changes: choose the reaction from the evidence

A copper wire can stay unchanged in dilute acid while magnesium fizzes. A battery can then force a reaction that would not happen by mixing alone. Use the **reactivity series** , the particles present and the reaction conditions to explain the difference. Sheets 4.1–4.3 contain original practice and authored marking guidance; Higher Tier (HT) demands are labelled.

Reactivity of metals (4.4.1.1–4.4.1.4)

Metal atoms form positive ions. A more reactive metal has a greater tendency to do this. Learn potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper in decreasing reactivity. Carbon and hydrogen are useful reference points; they are not metals.



Carbon and hydrogen are non-metals. Compare metals under matching conditions.

At room temperature, potassium, sodium and lithium react with water to form a soluble hydroxide and hydrogen. Sodium usually melts into a ball; lithium floats and fizzes; potassium may burn with a lilac flame. Calcium fizzes and forms calcium hydroxide; the mixture can become cloudy. Magnesium reacts extremely slowly with cold water; zinc, iron and copper show no observable reaction. Compare matching conditions and clean, similar-sized samples. Steam reactions are outside this section.

With dilute hydrochloric or sulfuric acid, magnesium, zinc and iron make a **salt** and hydrogen; copper does not. The more reactive metals above magnesium react too vigorously for an ordinary classroom acid comparison. Wear eye protection; keep hydrogen away from flames during collection. The specified acid–metal chemistry uses Mg, Zn and Fe with these two acids; do not apply its hydrogen rule to nitric acid.

Displacement occurs when a more reactive metal replaces a less reactive metal from a compound. For zinc in copper(II) sulfate:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Copper forms and the blue solution fades as copper ions are removed. A result can locate a metal between two known metals; one test alone cannot determine its whole rank.

Metals gain oxygen to make oxides: $2Mg + O_2 \rightarrow 2MgO$. This is **oxidation** . Loss of oxygen is **reduction** . Gold can occur as the metal itself because it is unreactive. Most metals occur as compounds. Carbon can reduce oxides of metals below it; a metal above carbon needs another method such as electrolysis. Detailed industrial extraction processes are not required here; use any extra process information supplied in a question.

Worked oxygen-transfer example. In $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$, iron oxide loses oxygen and is reduced. Carbon gains oxygen and is oxidised. The equation conserves four iron atoms, six oxygen atoms and three carbon atoms.

HT: electron transfer. Oxidation is loss of electrons; reduction is gain. For zinc/copper displacement, omit unchanged sulfate ions:

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

$$Zn \rightarrow Zn^{2+} + 2e^- \qquad Cu^{2+} + 2e^- \rightarrow Cu$$

Zinc atoms are oxidised; copper ions, not copper atoms, are reduced. For $Mg + 2HCl \rightarrow MgCl_2 + H_2$, Mg loses two electrons and two H^+ ions gain them; chloride is unchanged.

Reactions of acids and making salts (4.4.2)

An acid in water produces H^+ ions. An alkali is a soluble base whose aqueous solution contains OH^- ions; it does not donate electrons to neutralise an acid. **Neutralisation** between acid and base makes salt and water. Carbonates also produce carbon dioxide. A metal reaction instead produces hydrogen.

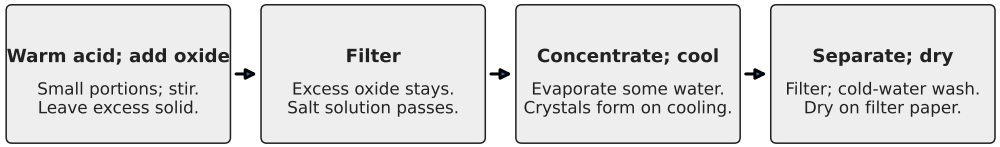
- Hydrochloric acid makes chlorides; nitric acid makes nitrates; sulfuric acid makes sulfates.
- The positive ion comes from the metal, base or carbonate. Balance charges before balancing the equation.
- Example: Ca^{2+} and NO_3^- make $Ca(NO_3)_2$, not $CaNO_3$.

$$CaCO_3 + 2HNO_3 \rightarrow Ca(NO_3)_2 + CO_2 + H_2O$$

- $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ is the ionic neutralisation equation.

Required practical 1: pure dry copper(II) sulfate crystals. Gently warm dilute sulfuric acid, add copper(II) oxide in small portions with stirring until some black solid remains, and filter. Excess insoluble solid shows that the acid has reacted; filtering removes that excess. Heat the blue filtrate gently in a water bath or with an electric heater to concentrate it, without boiling it dry. Leave it to cool and crystallise. Filter the crystals, wash with a little cold distilled water and dry between filter papers. Washing removes solution from the surface; use little cold water to limit crystal loss. Wear eye protection and handle hot apparatus carefully. This written method does not replace supervised practical work.

Pure dry copper(II) sulfate: choose each step by what is soluble

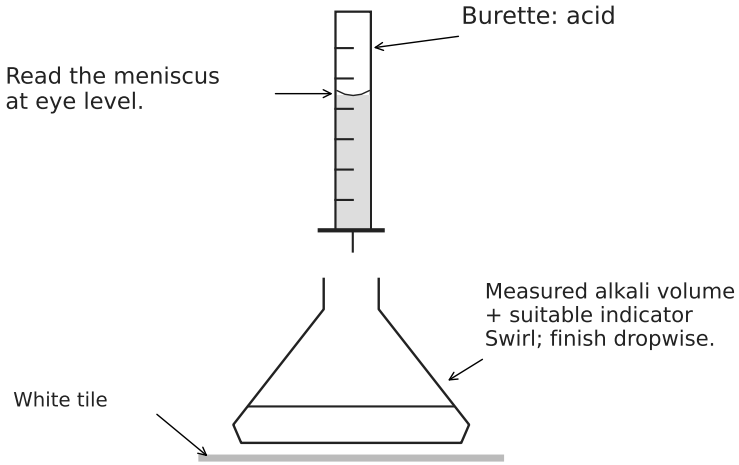


Do not try to filter dissolved acid; do not boil the salt solution dry.

pH scale. At this level, pH runs from 0 to 14: below 7 acidic, 7 neutral, above 7 alkaline. Use universal indicator with its colour chart for an approximate pH, or a suitable pH probe for a reading. Add indicator to a small sample, not the stock bottle. A single narrow-range indicator detects a suitable **titration** endpoint but does not measure the full pH range. Diluting an acid lowers H^+ concentration and brings pH towards 7; it does not make the acid an alkali.

Required practical 2, chemistry only: titration. Rinse a pipette with the alkali, use a pipette filler to transfer a fixed volume into a conical flask, and add a few drops of a suitable indicator. Rinse the burette with the acid; fill it and remove the jet air bubble. Read the bottom of the meniscus at eye level and record initial and final readings. Swirl the flask and add acid slowly, dropwise near the endpoint, over a white tile. The titre is final minus initial reading. Use a rough trial, then repeat accurate runs and average **concordant titres** using the stated tolerance; exclude the rough result. Adding distilled water to the flask washes down splashes without changing the alkali amount. Actual practical work remains required.

Titre = final burette reading − initial burette reading



Worked HT calculation. 20.0 cm^3 of 0.100 mol/dm^3 NaOH reacts with 25.0 cm^3 sulfuric acid. Find the acid concentration and its mass concentration; $M(H_2SO_4) = 98$ g/mol.

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

$$V(\text{NaOH}) = \frac{20.0}{1000}\ \text{dm}^3 = 0.0200\ \text{dm}^3$$

$$n(\text{NaOH}) = cV = (0.100\ \text{mol/dm}^3)(0.0200\ \text{dm}^3) = 0.00200\ \text{mol}$$

$$n(H_2SO_4) = \frac{n(\text{NaOH})}{2} = \frac{0.00200\ \text{mol}}{2} = 0.00100\ \text{mol}$$

$$c(H_2SO_4) = \frac{n}{V} = \frac{0.00100\ \text{mol}}{0.0250\ \text{dm}^3} = 0.0400\ \text{mol/dm}^3$$

$$c_m = cM = (0.0400\ \text{mol/dm}^3)(98\ \text{g/mol}) = 3.92\ \text{g/dm}^3$$

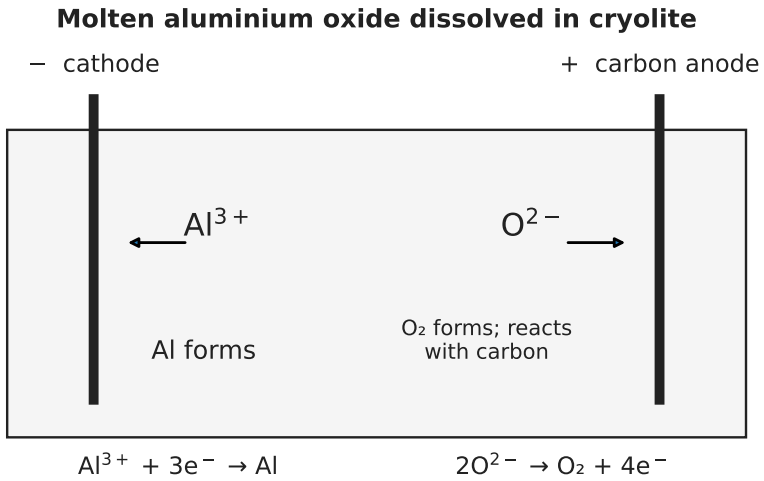
HT: strong and weak acids. A **strong acid** is completely ionised in water; hydrochloric, nitric and sulfuric acids are specified examples. A **weak acid** is only partially ionised; ethanoic, citric and carbonic acids are examples. At a given concentration, a stronger acid has lower pH. Concentrated/dilute describes amount of acid per solution volume, whereas strong/weak describes degree of ionisation. Diluting a strong acid does not turn it into a weak acid. A drop of one whole pH unit means ten times the H^+ concentration; pH 2 has 1000 times the H^+ concentration at pH 5. Neutrality concerns the balance of H^+ and OH^-, not the absence of ions.

Electrolysis (4.4.3.1–4.4.3.5)

An **electrolyte** conducts when its ions can move: an ionic compound may be molten or dissolved in water. Ions cannot carry charge through a solid lattice because they cannot move from place to place. **Electrolysis** uses a current to discharge ions at electrodes and produce elements. Cations move to the negative **cathode** and gain electrons; anions move to the positive **anode** and lose electrons. Ions carry charge in the liquid; electrons carry charge through the external metal circuit.

Molten binary compounds, inert electrodes. Lead bromide gives lead at the cathode and bromine at the anode. Molten zinc chloride gives zinc and chlorine. There is no water to compete in a molten salt.

Aluminium extraction. Aluminium is above carbon. Electrolysis uses much electrical energy for the current and heating. Aluminium oxide dissolves in molten cryolite; the mixture has a lower melting point than pure aluminium oxide, reducing heating demand. Oxygen formed at the positive carbon electrode reacts with it to make carbon dioxide, so this electrode must be replaced. The diagram applies to this molten mixture, not an aqueous aluminium salt.



Mobile ions carry charge in the melt; electrons move through the external circuit.

Aqueous solutions with inert electrodes. Water supplies H^+ and OH^- as additional possibilities. In the specified GCSE rule, a metal below hydrogen deposits at the cathode; if the metal is above hydrogen, hydrogen forms. At the anode, a halide gives its halogen; otherwise oxygen forms. Concentration and electrode material can matter in real systems; use the conditions supplied and do not transfer an inert-electrode prediction to reactive copper electrodes without extra information.

Electrolyte	Cathode	Anode
Aqueous NaCl	hydrogen	chlorine
Aqueous $CuCl_2$	copper	chlorine
Aqueous Na_2SO_4	hydrogen	oxygen
Aqueous $CuSO_4$	copper	oxygen

Required practical 3. Develop a hypothesis for each solution before testing with inert electrodes. Use labelled separate samples, the same electrode material/spacing, suitable low-voltage DC and matching test time. Observe each electrode separately. Test hydrogen with a lighted splint on a small collected sample (squeaky pop); oxygen relights a glowing splint; chlorine bleaches damp litmus paper. A copper deposit is pink-brown. Wear eye protection, avoid skin contact, and use only small gas quantities with good ventilation and teacher supervision; never smell chlorine. Wash electrodes between solutions to reduce contamination. Compare actual evidence with the hypothesis; absence of visible gas is not proof of no reaction. Written prediction does not replace the practical.

HT: half equation . First balance atoms, then charge using electrons; electrons are on the left for reduction and right for oxidation.

$$Al^{3+} + 3e^- \rightarrow Al \qquad 2O^{2-} \rightarrow O_2 + 4e^-$$

$$Cu^{2+} + 2e^- \rightarrow Cu \qquad 2Cl^- \rightarrow Cl_2 + 2e^-$$

$$2H^+ + 2e^- \rightarrow H_2 \qquad 4OH^- \rightarrow O_2 + 2H_2O + 4e^-$$

In the oxide equation the reactant is two oxide ions, not peroxide or an oxygen molecule with a charge. Check both oxygen atoms and total −4 charge. In the hydroxide equation four oxygen and four hydrogen atoms occur on each side, with total −4 charge on each side.

Checklist before you call this topic done

- Use observed reactions to order unknown metals; distinguish oxygen-transfer and HT electron-transfer redox.
- Choose reagents and a separation method for a named pure dry salt; deduce its formula from charges.
- Explain pH measurement, dilution, strength and concentration; carry out chemistry titration methods and HT calculations with the balanced ratio.
- Predict molten/aqueous products, explain aluminium extraction, and test an RP3 hypothesis; HT students balance both half-equations.
- Attempt unfamiliar tasks in sheets 4.1–4.3 before viewing the marking guidance. Practical participation and authentic-source integration are separate checks.