The Reactivity Series

Potassium	most reactive
Sodium	↑
Calcium	
Magnesium	
Aluminium	
Carbon	
Zinc	
Iron	
Tin	
Lead	
Hydrogen	
Copper	
Silver	
Gold	↓
Platinum	least reactive

When metals react with other substances the metal atoms form positive ions.

The reactivity of a metal is related to its tendency to form positive ions. More reactive metals form positive ions faster.

Metals can be arranged in order of their reactivity in a "reactivity series" based on how they react with water, oxygen, dilute acids and each other. You must learn this order.

A more reactive metal can **displace** a less reactive metal from a compound. E.g. Ca + ZnO —> CaO + Zn (The calcium **displaces** the zinc)

Reactions of metal with acids

Metal + acid → salt + hydrogen (tip: MASH)
E.g. zinc + hydrochloric acid → zinc chloride + hydrogen

Reactions of metal with oxygen

Metal + oxygen → metal oxide
E.g. iron + oxygen → iron oxide
This is an oxidation reaction because the metals gains oxygen.

Reactions of metal oxides with acids

Metal oxide + acid → salt + water
E.g. magnesium oxide + sulfuric acid → magnesium sulfate + water

Reactions of metal carbonates with acids

Metal carbonate + acid → salt + water + carbon dioxide
E.g. tin carbonate + hydrochloric acid → tin chloride + water + carbon dioxide

Oxidation and Reduction (Higher Tier only)

When a metal forms a bond with a non-metal element it loses its outershell electron(s). The metal is oxidised.
When a metal is in a compound and reacts to form an element it gains electron(s). The metal is reduced.
Metal oxidation: zinc + hydrochloric acid → zinc chloride + hydrogen
Metal reduction: copper oxide + carbon → copper + carbon dioxide
OILRIG: **o**xidation **i**s **l**oss, **r**eduction **i**s **g**ain

Metal Extraction (a specific displacement reaction)

Metals can be split into three groups, based on how easy to extract them it is.

1. The most unreactive metals (e.g. gold) are found **native** (as an unreacted element) in the Earth's crust.
2. Metals **less reactive than carbon** can be extracted from compounds by **reduction** with carbon. This is a special example of a displacement reaction

   E.g. iron oxide + carbon —> iron + carbon dioxide
3. Metals **more reactive than carbon** may require **electrolysis (Chapter 6)** to extract them. This is **expensive** and needs a lot of **energy**.

Salts

Salts are ionic compounds. They are named according to the acid and the metal. The metal can either exist as a compound or as a pure element. The metal must be more reactive than hydrogen in order to react with an acid.

1. **Hydrochloric acid** (HCl) reacts to make **chloride** salts

   E.g. iron + hydrochloric acid —> iron chloride + hydrogen
2. **Sulfuric acid** (H_2SO_4) reacts to make **sulfate** salts

   E.g. zinc + sulfuric acid —> zinc sulfate + hydrogen
3. **Nitric acid** (HNO_3) reacts to make **nitrate** salts

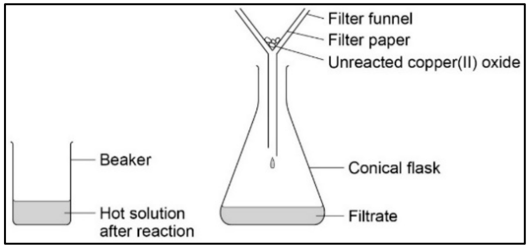
   E.g. magnesium + nitric acid —> magnesium nitrate + hydrogen

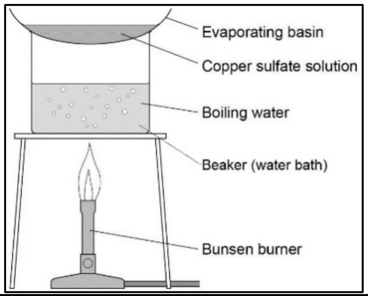
Making a named soluble salt (required practical)

Soluble salts are those that can dissolve. Soluble salts can be made from acids by reacting them with solid insoluble substances, such as metals, metal oxides, hydroxides or carbonates.

1) The solid metal compound is added to the acid until no more reacts
2) The **excess solid** is filtered off to produce a solution of the salt.
3) The salt solution is heated then then left to crystallise to produce solid salt.

An excess of the solid MUST be added to make sure no acid remains to form an impurity in the salt product.





Neutralisation

A **neutralisation** reaction occurs between an acid and a base. **Alkali's** are a type of base which are soluble in water. Not all bases are soluble in water e.g. ALL metal oxides and SOME metal hydroxides.

A neutralisation reaction between an acid and a base produces a salt and water.

The water is formed from the H^+ ions in the acid and the OH^- ions in the base.

E.g. sodium hydroxide + nitric acid —> sodium nitrate + water

It can also be partially represented by the equation: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O\ (l)$

Strong and weak acids (Higher only)

Acids all release H^+ ions when they are dissolved in water. This is called being "in an aqueous solution".

Some acids like hydrochloric acid (HCl), sulfuric acid and nitric acid fully ionise to release H^+ ions extremely easily. These are called **strong acids**.

Some acids like ethanoic acid (CH_3COOH) are only partially ionised in aqueous solution. This means that some of the particles will split up (to make CH_3COO^- and H^+) but some of them will remain as CH_3COOH. These are called **weak acids**.

Other examples of weak acids are citric acid and carbonic acid.

Determining concentration with a titration (required practical)

1. Measure an exact volume of a known concentration of alkali to a conical flask using a volumetric pipette.
2. Add a few drops of pH indicator into the flask.
3. Fill a burette with acid of unknown concentration.
4. Add acid to alkali, swirling solution, until the end-point is reached (indicator remains a permanent colour)
5. Note the volume added and repeat until concordant titres are obtained.

Calculate the concentration of unknown acid

a. Calculate mean titre volume from concordant results only.
b. Calculate moles of alkali used, factor any mole ratios between acid and alkali then divide by the mean titre (in dm^3)

The pH scale

The pH scale is a measure of acidity, ranging from 0 (extremely acidic) to 14 (extremely alkaline).

Acidic solutions produce **hydrogen (H^+)** ions.

Alkaline solutions produce **hydroxide (OH^-)** ions.

A solution with pH 7 is neutral.

pH can be measured using **indicators** (e.g. universal indicator) or a **pH probe** (a detector attached to a computer).

An indicator is a chemical that changes colour in response to differences in pH

Universal indicator is **red** in acids, **purple** in alkalis and **green** in neutral solutions.

pH and acids (Higher only)

The pH scale is a logarithmic scale. As pH decreases by one unit, the hydrogen ion concentration of the solution increases by a factor of 10.

This means that a solution with pH 4 has 10 times more H^+ ions in it than a solution with pH 5, and 100 times more H^+ ions that a solution with pH 6.

Strong acid vs Concentrated acid (Higher only)

A strong acid (which fully ionises in water) is different to a concentrated acid (where lots of the acid has been dissolved in a small amount of solvent).

A weak acid will have a higher (less acidic pH) than a strong acid of the same concentration.

Ionic equations and half-equations (Higher only)

$Fe(s) + CuSO_4 \rightarrow FeSO_4 + Cu$

The ionic equation for the addition of copper sulfate to iron is below:

$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Half-equations can be used to show what happens to each reactant in terms of reactant.

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ (iron loses two electrons to form an iron ion)

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (copper gains two electrons to form a copper atom)