Acids, Bases and Salts

The properties of acids and bases, neutralisation, indicators and pH, strong and weak acids, acidic, basic and amphoteric oxides, and the preparation of soluble and insoluble salts.

7.1 The characteristic properties of acids and bases

- An **acid** is a substance that forms H^+ **ions** when it dissolves in water. Aqueous solutions of acids contain H^+ ions.
- An **alkali** is a **soluble base**. Aqueous solutions of alkalis contain OH^- **ions**.

> **An acid is a proton donor and a base is a proton acceptor. A hydrogen ion, H^+, is just a proton.**

Bases are oxides or hydroxides of metals. Most bases are **insoluble** in water. The bases that **dissolve** in water are the **alkalis**, such as sodium hydroxide, potassium hydroxide and calcium hydroxide (partially).

Reactions of acids

Acid reacts with	Products	Example
a **metal** (more reactive than hydrogen)	salt + hydrogen	$Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
a **base** (metal oxide or hydroxide)	salt + water	$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$
a **carbonate**	salt + water + carbon dioxide	$CaCO_3(s) + 2\,HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

- A reaction with a metal gives off hydrogen: it burns with a squeaky pop.
- A reaction with a carbonate **fizzes** and the gas turns **limewater milky** (carbon dioxide).

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Reactions of bases

Base reacts with	Products	Example
an **acid**	salt + water	$\mathrm{NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)}$
an **ammonium salt** (on warming)	salt + water + ammonia	$\mathrm{NaOH(aq) + NH_4Cl(s) \longrightarrow NaCl(aq) + H_2O(l) + NH_3(g)}$

Ammonia is a gas with a sharp smell that turns **damp red litmus paper blue**.

Neutralisation

A **neutralisation** reaction occurs between an **acid and a base** to make a salt and water. When an acid reacts with an alkali, the ionic equation is the same for every pair:

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

Indicators and pH

Indicator	In acid	In alkali
litmus	red	blue
thymolphthalein	colourless	blue
methyl orange	red	yellow

Universal indicator (a mixture) gives a range of colours that match the **pH** number, and so it shows **how acidic or alkaline** a solution is. It is often used as paper or as a solution.

- **pH** < 7**:** acid. The lower the pH, the **higher the concentration of** $\mathrm{H^+}$ **ions** and the more acidic the solution.
- **pH** $= 7$**:** neutral (pure water).
- **pH** > 7**:** alkaline. The higher the pH, the more alkaline.

Strong and weak acids

> **A strong acid is completely dissociated (fully ionised) in aqueous solution. A weak acid is partially dissociated.**

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Strong acids include hydrochloric acid, nitric acid and sulfuric acid. All of their molecules split into ions, shown with a single arrow:

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

$$H_2SO_4(aq) \longrightarrow 2\,H^+(aq) + SO_4{}^{2-}(aq)$$

Weak acids include the carboxylic acids such as ethanoic acid. Only a small fraction of the molecules split into ions, and an equilibrium is set up, shown with a reversible arrow:

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

At the **same concentration**, a strong acid has a **lower pH** (more H^+ ions), a **faster reaction with a metal**, and conducts electricity better than a weak acid.

NOTE

Strong and **weak** describe how much an acid **dissociates**. **Concentrated** and **dilute** describe how much acid is **dissolved** in the solution. A dilute solution of a strong acid can have a higher pH than a concentrated solution of a weak acid, but they are different ideas.

7.2 Oxides

Oxides are classified by how they react with acids and bases.

Type	Which elements	Reacts with	Examples
acidic oxide	**non-metals**	bases and alkalis	SO_2, CO_2
basic oxide	**metals**	acids	CuO, CaO
amphoteric oxide	between metals and non-metals	**both** acids and bases	Al_2O_3, ZnO

An amphoteric oxide is an oxide that reacts with both acids and bases to produce a salt and water.

$$CaO(s) + 2\,HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) \quad \text{(basic oxide with an acid)}$$

$$CO_2(g) + Ca(OH)_2(aq) \longrightarrow CaCO_3(s) + H_2O(l) \quad \text{(acidic oxide with an alkali)}$$

$$ZnO(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + H_2O(l) \quad \text{(amphoteric oxide with an acid)}$$

Zinc oxide also reacts with sodium hydroxide solution to give a salt and water, which shows that it is amphoteric. Aluminium oxide does the same.

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The oxides of **metals** are usually **basic**, and the oxides of **non-metals** are usually **acidic**. Aluminium and zinc are in between, so their oxides are amphoteric.

7.3 Preparation of salts

A **salt** is a compound made when the H^+ of an acid is replaced by a metal ion or by NH_4^+. The first part of the name comes from the base or metal and the second part from the acid.

Acid	Salts it forms
hydrochloric acid, HCl	chlorides
sulfuric acid, H_2SO_4	sulfates
nitric acid, HNO_3	nitrates
ethanoic acid, CH_3COOH	ethanoates

Solubility rules

Soluble	Insoluble (exceptions)
all **sodium, potassium and ammonium** salts	
all **nitrates**	
chlorides	except **silver** and **lead** chlorides
sulfates	except **barium**, **calcium** and **lead** sulfates
	carbonates, except sodium, potassium and ammonium
	hydroxides, except sodium, potassium, ammonium and calcium (partly soluble)

Making a soluble salt

The method depends on whether the other reactant is soluble.

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Acid plus	Method
an alkali (such as sodium hydroxide)	**titration**: use a burette and an indicator to find the exact volume of acid that neutralises a measured volume of alkali. Then repeat **without the indicator**, and evaporate the solution to get the salt
an excess of a metal	add the metal to the acid until no more reacts (no more bubbles), then filter off the extra metal
an excess of an insoluble base (oxide or hydroxide)	warm the acid and add the solid until no more dissolves, then filter
an excess of an insoluble carbonate	add the carbonate to the acid until the **fizzing stops** (no more carbon dioxide), then filter

The steps (after the reaction using an excess of a solid):

1. **Filter** to remove the excess solid. The **filtrate** is the salt solution.
2. **Heat** the filtrate in an evaporating basin to **evaporate** some of the water, until the solution is **saturated** (a drop on a glass rod crystallises when cooled).
3. Leave it to **cool and crystallise**.
4. Remove the crystals by filtering, **wash** them with a little cold distilled water, and **dry** them between filter papers.

> NOTE
>
> An **excess** of the insoluble solid is used so that **all the acid** is used up. The excess is easy to remove by filtering, but extra soluble reactant would stay in the solution and make the salt impure.

> TRY IT
>
> Describe how to make crystals of copper(II) sulfate from copper(II) oxide and dilute sulfuric acid.

Warm the dilute sulfuric acid and add copper(II) oxide a little at a time, stirring, until no more dissolves (the oxide is in excess). **Filter** to remove the excess copper(II) oxide. **Heat the filtrate** to evaporate some of the water until it is saturated, then leave it to **cool and crystallise**. Filter off the blue crystals, wash with a little cold distilled water, and dry between filter papers.

$$\mathrm{CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)}$$

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Making an insoluble salt by precipitation

An **insoluble** salt is made by mixing two **soluble** compounds that contain its ions. The insoluble salt forms as a **precipitate**.

> **TRY IT**
>
> How can barium sulfate, an insoluble salt, be prepared?

Mix a solution of a soluble barium salt, such as barium chloride, with a solution of a soluble sulfate, such as sodium sulfate. A white precipitate of barium sulfate forms. **Filter** it, **wash** it with distilled water to remove the soluble impurities, and **dry** it.

$$Ba^{2+}(aq) + SO_4{}^{2-}(aq) \longrightarrow BaSO_4(s)$$

Hydrated and anhydrous salts

> **A hydrated substance is chemically combined with water. An anhydrous substance contains no water. The water of crystallisation is the water molecules in hydrated crystals.**

Examples: hydrated copper(II) sulfate, $CuSO_4 \cdot 5\,H_2O$ (blue), and hydrated cobalt(II) chloride, $CoCl_2 \cdot 6\,H_2O$ (pink). The "5" and the "6" show the number of moles of water for each mole of salt.

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