

Acids, bases and salts

IGCSE Chemistry 0620 · Topic 7 revision notes

Everything from the 7 lessons of this topic on a few pages: the key idea of each lesson, the diagram to remember, the points to learn and the traps examiners set. Read it, cover it, test yourself.

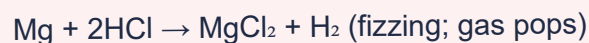
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| 1. <u>Properties of acids and bases</u> | 4. <u>Acidic, basic and amphoteric oxides</u> |
| 2. <u>pH, indicators and neutralisation</u> | 5. <u>Preparing soluble salts</u> |
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1 Properties of acids and bases

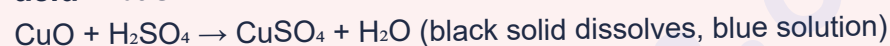
Syllabus 7.1

Acids turn blue litmus red, turn methyl orange red and leave thymolphthalein colourless. They react with **metals** to give a salt and hydrogen, with **bases** to give a salt and water, and with **carbonates** to give a salt, water and carbon dioxide. **Bases** are oxides or hydroxides of metals; soluble bases are **alkalis**. Alkalis turn red litmus blue, methyl orange yellow and thymolphthalein blue, and release ammonia from ammonium salts. Acids contain H^+ ions and alkalis contain OH^- ions.

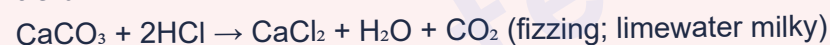
acid + metal $\rightarrow$ salt + hydrogen



acid + base $\rightarrow$ salt + water



acid + carbonate $\rightarrow$ salt + water + carbon dioxide



alkali + ammonium salt $\rightarrow$ salt + water + ammonia



Fig. 1: the four reaction patterns to learn

Four patterns, one salt.

- ✓ **Metals:** salt + hydrogen.
- ✓ **Bases:** salt + water.
- ✓ **Carbonates:** salt + water + carbon dioxide.
- ✓ **Indicators:** litmus red/blue, methyl orange red/yellow, thymolphthalein colourless/blue.

Watch out!

Metal + acid gives **hydrogen**; carbonate + acid gives **carbon dioxide**. Mixing these up is the commonest slip. Also, a base is not always an alkali: say "alkali" only for a base that dissolves in water.

2 pH, indicators and neutralisation

Syllabus 7.1

The **pH scale** runs from 1 to 14. pH 7 is neutral; below 7 is acidic and above 7 is alkaline. **Universal indicator** gives a different colour for each pH, from red (strongly acidic) through green (neutral) to purple (strongly alkaline). The lower the pH, the higher the concentration of H^+ ions. In **neutralisation**, H^+ and OH^- ions react to form water, so the pH moves towards 7.

universal indicator colours on the pH scale

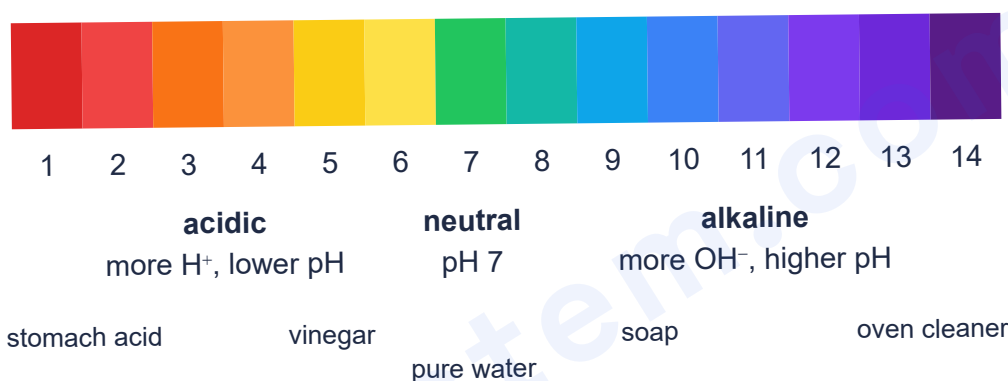


Fig. 1: the pH scale and universal indicator

Lower number, more acid.

- ✓ **pH 7** neutral; below acidic; above alkaline.
- ✓ **Universal indicator:** red, orange, yellow, green, blue, purple.
- ✓ **Lower pH** = higher H^+ concentration.
- ✓ **Neutralisation** moves the pH towards 7.

Watch out!

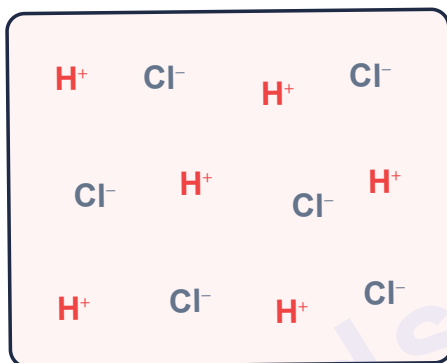
A **lower** pH means a **higher** concentration of H^+ ions, not a lower one. And pH 7 means neutral only if the solution contains no excess acid or alkali; "more acidic" means a smaller pH number.

3 Acids as proton donors, strong and weak acids

Syllabus 7.1

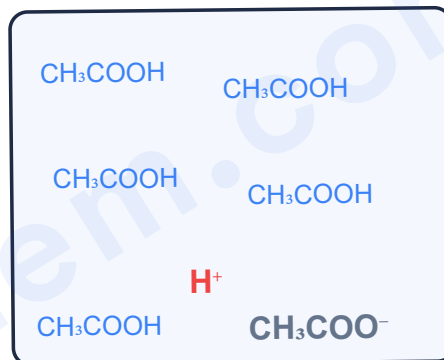
Extended. An **acid** is a **proton donor** and a **base** is a **proton acceptor**, because an H^+ ion is a proton. A **strong acid** is completely dissociated in water: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$. A **weak acid** is only partially dissociated: $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$. At the same concentration a weak acid has fewer H^+ ions, a higher pH and reacts more slowly.

strong: hydrochloric acid



completely dissociated
 $HCl \rightarrow H^+ + Cl^-$

weak: ethanoic acid



partially dissociated
 $CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$

Fig. 1: complete and partial dissociation

Donate, accept, dissociate.

- ✓ **Acid:** proton donor. **Base:** proton acceptor.
- ✓ **Strong:** completely dissociated ($\rightarrow$).
- ✓ **Weak:** partially dissociated ($\rightleftharpoons$).
- ✓ **Strength $\neq$ concentration.**

Watch out!

Strong and **concentrated** are different ideas. Strength is about how completely the acid dissociates; concentration is about how much acid is dissolved per dm^3 . A dilute solution of a strong acid and a concentrated solution of a weak acid are both possible.

4 Acidic, basic and amphoteric oxides

Syllabus 7.2

Basic oxides are oxides of metals, such as calcium oxide and copper(II) oxide; they react with acids to form a salt and water. **Acidic oxides** are oxides of non-metals, such as carbon dioxide and sulfur dioxide; they react with bases to form a salt and water, and many dissolve in water to give acidic solutions. **Extended: amphoteric oxides**, such as aluminium oxide and zinc oxide, react with both acids and bases.

basic

most metal oxides

CaO, MgO, CuO,
Na₂O

react with acids
to form a salt
and water

amphoteric

some metal oxides

Al₂O₃, ZnO

react with acids
AND with bases
(Extended)

acidic

most non-metal
oxides

CO₂, SO₂, P₄O₁₀

react with bases
to form a salt
and water

metal on the left of the Periodic Table: basic; non-metal on the right: acidic

Fig. 1: three kinds of oxide

Metal basic, non-metal acidic.

- ✓ **Basic oxides** (metals) react with acids.
- ✓ **Acidic oxides** (non-metals) react with bases.
- ✓ **Amphoteric** (Al₂O₃, ZnO) react with both.
- ✓ **Across a period:** basic → amphoteric → acidic.

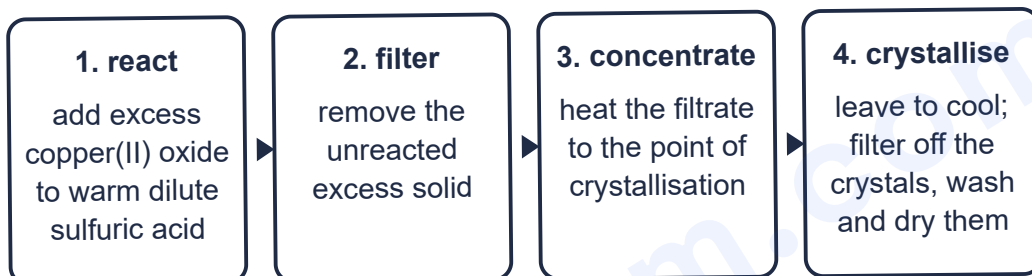
Watch out!

Not every metal oxide is simply basic: aluminium and zinc oxides are amphoteric. And "acidic oxide" means it reacts with bases; it does not contain hydrogen ions itself. Silicon dioxide is acidic even though it does not dissolve in water.

5 Preparing soluble salts

Syllabus 7.3

A **salt** forms when the hydrogen of an acid is replaced by a metal or ammonium ion. To make a soluble salt from an insoluble solid: add **excess** solid (metal, base or carbonate) to warm acid, **filter** off the excess, **heat** the filtrate to the point of crystallisation, leave to **cool**, then filter, wash and dry the crystals. For sodium, potassium and ammonium salts, react an acid with an alkali using a **titration**, then repeat without the indicator and crystallise.



excess solid makes sure all the acid is used up,
so the crystals are not contaminated with acid

Fig. 1: making copper(II) sulfate crystals by the excess solid method

Excess, filter, crystallise.

- ✓ **Insoluble reactant:** excess solid method.
- ✓ **Na, K, NH₄ salts:** titration method.
- ✓ **Point of crystallisation,** then cool; never to dryness.
- ✓ **Filter, wash, dry** the crystals at the end.

Watch out!

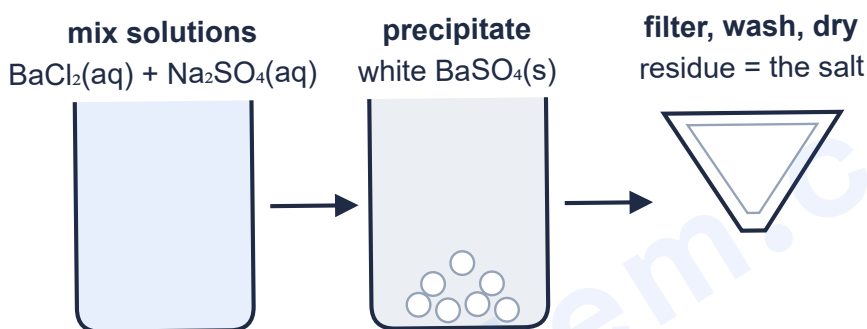
Filter **before** crystallising, to remove the excess solid; then filter again **after** crystallising, to separate the crystals from the remaining solution. And the salt is in the **filtrate** in the first filtration, not in the residue.

6 Insoluble salts and solubility rules

Syllabus 7.3

General rules: all **sodium, potassium and ammonium** salts are soluble; all **nitrates** are soluble; **chlorides** are soluble except silver and lead(II) chlorides; **sulfates** are soluble except barium, calcium and lead(II) sulfates; **carbonates** and **hydroxides** are insoluble except those of sodium, potassium and ammonium (calcium hydroxide is slightly soluble).

Extended: an insoluble salt is made by **precipitation**: mix two solutions that contain its ions, then filter, wash and dry the precipitate.



wash with distilled water to remove the soluble sodium chloride

Fig. 1: making barium sulfate by precipitation

Know the exceptions.

- ✓ **Always soluble:** Na, K, NH_4 compounds and all nitrates.
- ✓ **Chlorides:** not Ag, Pb. **Sulfates:** not Ba, Ca, Pb.
- ✓ **Carbonates, hydroxides:** insoluble except Na, K, NH_4 .
- ✓ **Precipitation:** mix, filter, wash residue, dry.

Watch out!

For an insoluble salt the product you want is the **residue** on the filter paper, not the filtrate: the opposite of the soluble salt method. Nitrates are useful starting materials because they are always soluble; for example, use lead(II) **nitrate**, not lead(II) chloride, as the source of lead ions.

7 Hydrated and anhydrous salts

Syllabus 7.3

A **hydrated** substance is chemically combined with water; an **anhydrous** substance contains no water. **Water of crystallisation** is the water molecules present in the crystals of a hydrated salt, as in $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. Heating a hydrated salt drives off this water: blue hydrated copper(II) sulfate becomes white anhydrous copper(II) sulfate. **Extended:** the number of water molecules, x , can be found from the masses before and after heating.



heating drives off water; adding water turns it blue and hot
water of crystallisation: water molecules
that are part of the crystal structure

2.50 g hydrated $\rightarrow$ 1.60 g anhydrous: 0.90 g of water lost

Fig. 1: hydrated and anhydrous copper(II) sulfate

Mass lost is water.

- ✓ **Hydrated:** contains water of crystallisation. **Anhydrous:** none.
- ✓ **CuSO_4 :** blue hydrated, white anhydrous.
- ✓ **CoCl_2 :** pink hydrated, blue anhydrous.
- ✓ **x** = mol of water $\div$ mol of anhydrous salt.

Watch out!

Hydrated crystals look and feel **dry**: the water of crystallisation is locked inside the crystal structure, not wetting the surface. And removing it is a chemical change, shown by the colour change, not simply drying.