Cambridge (CIE) IGCSE Chemistry

Your notes

The Characteristic Properties of Acids & Bases

Contents

Properties of acids

- Acids have pH values of below 7, have a sour taste (when edible) and are corrosive
- Acids are substances that can neutralise a base, forming a salt and water
- When **acids** are added to water, they form positively charged **hydrogen ions** (H^+)
- The presence of H^+ ions is what makes a solution acidic

Example: Hydrochloric acid

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

Acids and metals

- Only metals **above hydrogen** in the reactivity series will react with dilute acids.
- When acids react with metals they form a salt and hydrogen gas:

acid + metal → salt + hydrogen

- The name of the salt is related to the name of the acid used, as it depends on the **anion** within the acid

Examples of the names of salts from specific acids and metals

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and hydrogen	$Mg + 2HCl \rightarrow MgCl_2 + H_2$
Sulfuric acid	Magnesium sulfate and hydrogen	$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$
Nitric acid	Magnesium nitrate and hydrogen	$Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2$

Acids with bases

- Metal **oxides** and metal **hydroxides** (**alkalis**) can act as **bases**
- When they react with acid, a **neutralisation** reaction occurs
- In all acid-base neutralisation reactions, **salt** and **water** are produced

acid + base → salt + water

Examples of reactions between acids and bases

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and water	$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$
Sulfuric acid	Magnesium sulfate and water	$MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$
Nitric acid	Magnesium nitrate and water	$Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$

Acids with metal carbonates

- Acids will react with metal carbonates to form the corresponding metal salt, carbon dioxide and water:

acid + metal carbonate → salt + carbon dioxide + water

Examples of reactions between acids and carbonates

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride, carbon dioxide and water	$MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$
Sulfuric acid	Magnesium sulfate, carbon dioxide and water	$MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + CO_2 + H_2O$
Nitric acid	Magnesium nitrate, carbon dioxide and water	$MgCO_3 + 2HNO_3 \rightarrow Mg(NO_3)_2 + CO_2 + H_2O$

What are indicators?

- Two colour indicators are used to distinguish between acids and alkalis
- Many plants contain substances that can act as indicators and the most common one is **litmus** which is extracted from lichens
- Synthetic indicators are organic compounds that are sensitive to changes in acidity and appear different colours in acids and alkalis
- **Thymolphthalein** and **methyl orange** are synthetic indicators frequently used in acid-alkali titrations

Two Colour Indicators Table

Indicator	Colour in acid	Colour in alkali

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Litmus	red	blue
Thymolphthalein	colourless	blue
Methyl orange	red	yellow

What are synthetic indicators?

- Synthetic indicators are used to show the endpoint in titrations as they have a very sharp change of colour when an acid has been neutralised by alkali and vice-versa
- Litmus is not suitable for titrations as the colour change is not sharp and it goes through a purple transition colour in neutral solutions making it difficult to determine an endpoint
- Litmus is very useful as an indicator paper and comes in red and blue versions, for dipping into solutions or testing gases

Properties of bases & alkalis

- Bases have pH values of above 7
- A base which is water-soluble is referred to as an **alkali**
- In basic (alkaline) conditions red litmus paper turns **blue**, methyl orange indicator turns **yellow** and thymolphthalein indicator turns **blue**
- Bases are substances which can neutralise an acid, forming a salt and water
- Bases are usually **oxides** or **hydroxides** of metals
- When **alkalis** are added to water, they form negative hydroxide ions (OH^-)
- The presence of the OH^- ions is what makes the aqueous solution an alkali

Example: Sodium hydroxide

$$NaOH\,(s) \rightarrow Na^+\,(aq) + OH^-\,(aq)$$

Bases and acids

- When bases react with an acid, a neutralisation reaction occurs
- Acids and bases react together in a neutralisation reaction and produce a salt and water:

Acid + base → salt + water

Examples of reaction between bases and acids

Acid	Name of products	Equation for reaction
Hydrochloric acid	Magnesium chloride and water	$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$

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Sulfuric acid	Magnesium sulfate and water	$MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$
Nitric acid	Magnesium nitrate and water	$Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$

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Alkalis and ammonium salts

- Ammonium salts undergo **decomposition** when warmed with an alkali
- Even though ammonia is itself a weak base, it is very **volatile** and can easily be displaced from the salt by another alkali
- A salt, water and ammonia are produced
    - For example:

$$\mathbf{NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3}$$

- This reaction is used as a chemical test to confirm the presence of the ammonium ion (NH_4^+)
- Alkali is added to the substance with gentle warming followed by the test for ammonia gas using damp red litmus paper
- The damp litmus paper will turn from **red** to **blue** if ammonia is present

Neutralisation reactions

What is a neutralisation reaction?

- Acids are a source of hydrogen ions, **H^+**
- Bases (or alkalis) are a sources of hydroxide ions, **OH^-**
- When they react together in a neutralisation reaction, the H^+ ions react with the OH^- ions to produce **water**
- This is the **net ionic equation** of all acid-base neutralisations and is what leads to a neutral solution, since water has a pH of 7:

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

- Not all reactions of acids are neutralisations:
  - For example when a metal reacts with an acid, although a salt is produced there is no water formed so it does not fit the definition of neutralisation

Examiner Tips and Tricks

Not all reactions of acids are neutralisations. For example, when a metal reacts with an acid, although a salt is produced there is no water formed so it does not fit the definition of neutralisation.

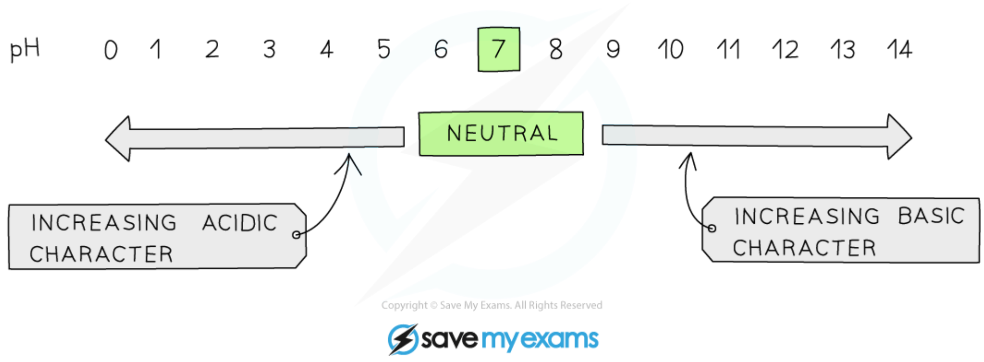
Hydrogen ion concentration & pH

What is the pH scale?

- The pH scale goes from 1 - 14
- All acids have pH values of **below** 7, all alkalis have pH values of **above** 7
- The **lower** the pH then the **more acidic** the solution is:
  - pH 0–2 = strong acid
    - Extremely acidic substances can have values of below 1
  - pH 3–6 = weak acid
- The **higher** the pH then the **more alkaline** the solution is:
  - pH 8–11 = weak alkali
  - pH 12–14 = strong alkali

- A solution of pH 7 is described as being **neutral**

The pH scale



The pH scale showing acidity, neutrality and alkalinity

The pH scale and hydrogen ions

- We have already seen that acids are substances that contain hydrogen ions in solution
- The more hydrogen ions the stronger the acid, but the lower the pH
- The higher the concentration of hydroxide ions in a solution the higher the pH
- So pH is a measure of the concentration of H^+ ions in solution, but they have an inverse relationship
- The pH scale is **logarithmic**, meaning that each change of 1 on the scale represents a change in concentration by a **factor** of **10**

Concentration of hydrogen ions and pH

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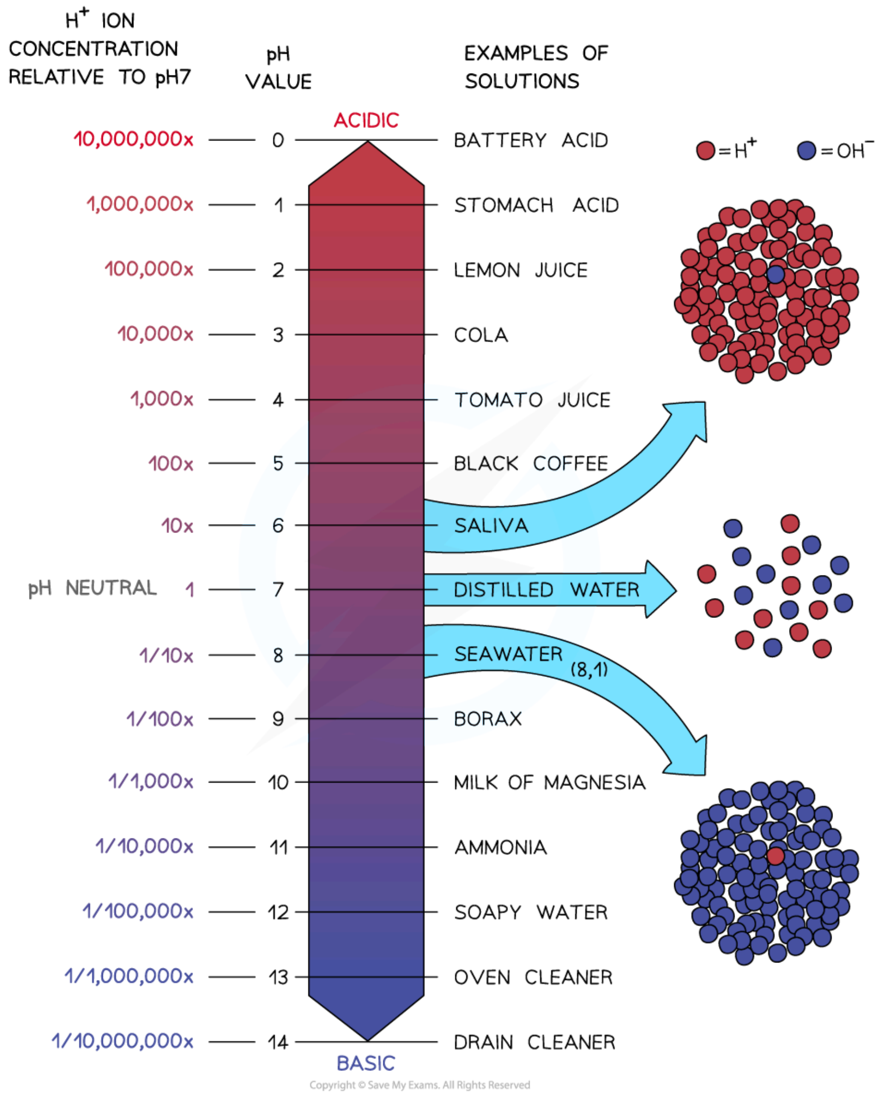

An acid with a pH of 3 has ten times the concentration of H^+ ions than an acid of pH 4. An acid with a pH of 2 has 10 x 10 = 100 times the concentration of H^+ ions than an acid with a pH of 4

Examiner Tips and Tricks

Acid strength is reflected in how many hydrogen ions are in solution. The more hydrogen ions the lower the pH and vice-versa.

How is universal indicator used?

- Universal indicator is a mixture of different **indicators** which is used to measure the pH
- A drop is added to the solution and the colour is matched with a colour chart which indicates the pH which matches specific colours

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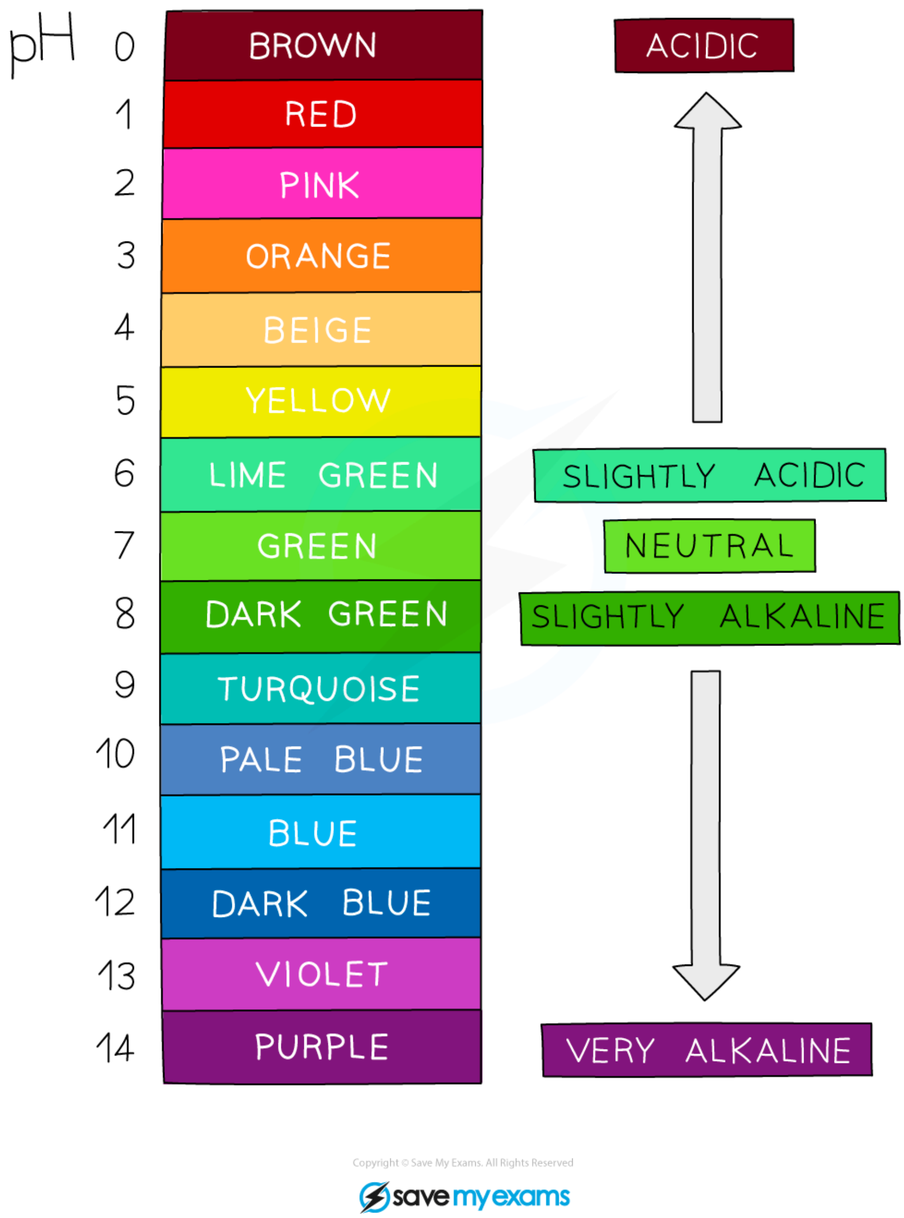

The pH scale with the Universal Indicator colours which can be used to determine the pH of a solution

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Proton transfer, strong & weak acids

Proton transfer

Extended tier only

- Acids are **proton donors** as they ionise in solution producing protons, which are H^+ ions
- These H^+ ions make the aqueous solution acidic
- Bases are **proton acceptors** as they accept the protons which are donated by the acid

Proton transfer between hydrochloric acid and water

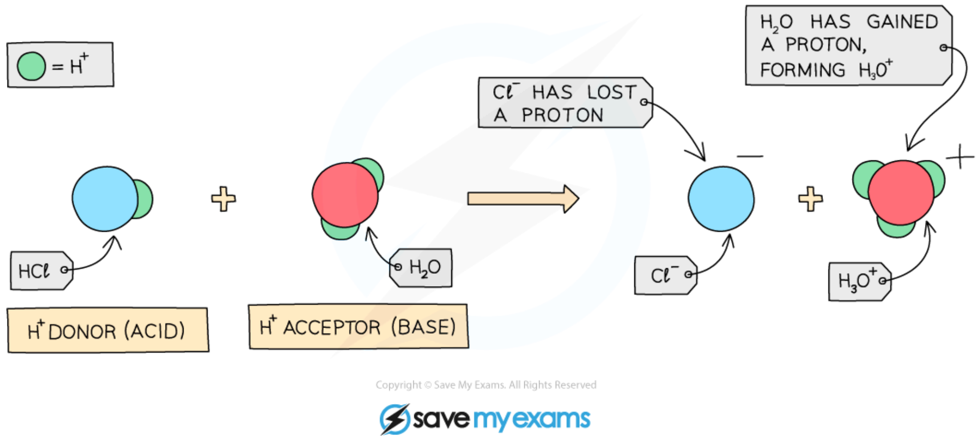

Diagram showing the role of acids and bases in the transfer of protons – here water acts as a base as it accepts a proton

What is a strong acid?

- Acids can be either **strong** or **weak**, depending on how many H^+ ions they produce when dissolved in water
- Strong acids **completely dissociate** (or ionise) in water, producing solutions of a **very low pH**
- Strong acids include HCl and H_2SO_4
- Example of a strong acid: hydrochloric acid

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

What is a weak acid?

- Weak acids **partially dissociate** (or ionise) in water and produce pH values which are closer to the **middle** of the pH scale, whilst still being below 7
- For weak acids, there is usually an equilibrium set-up between the molecules and their ions once they have been added to water
- Example of a weak acid: propanoic acid

$$CH_3CH_2COOH \rightleftharpoons H^+ + CH_3CH_2COO^-$$

- The equilibrium lies to the **left**, indicating a high concentration of intact acid molecules, with a low concentration of H^+ ions in the solution
- Another example of a weak acid is ethanoic acid which will react with alkalis such as sodium hydroxide to form ethanoate salts

ethanoic acid + sodium hydroxide → sodium ethanoate + water

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

Effect of concentration on strong and weak acids

- A concentrated solution of an acid is one that contains a **higher number** of acid **molecules** per dm^3 of solution
- It does not necessarily mean that the acid is strong though, as it may be made from a weak acid which does not **dissociate** completely
- For example a dilute solution of HCl will be more acidic than a concentrated solution of ethanoic acid, since most of the HCl molecules dissociate but very few of the CH_3COOH do

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Classifying oxides

What are oxides?

- Oxides are compounds made from one or more atoms of oxygen combined with one other element
- Examples of oxides include: MgO, ZnO, K_2O, CO_2, SO_2, H_2O
- Oxides can be classified based on their acid-base characteristics

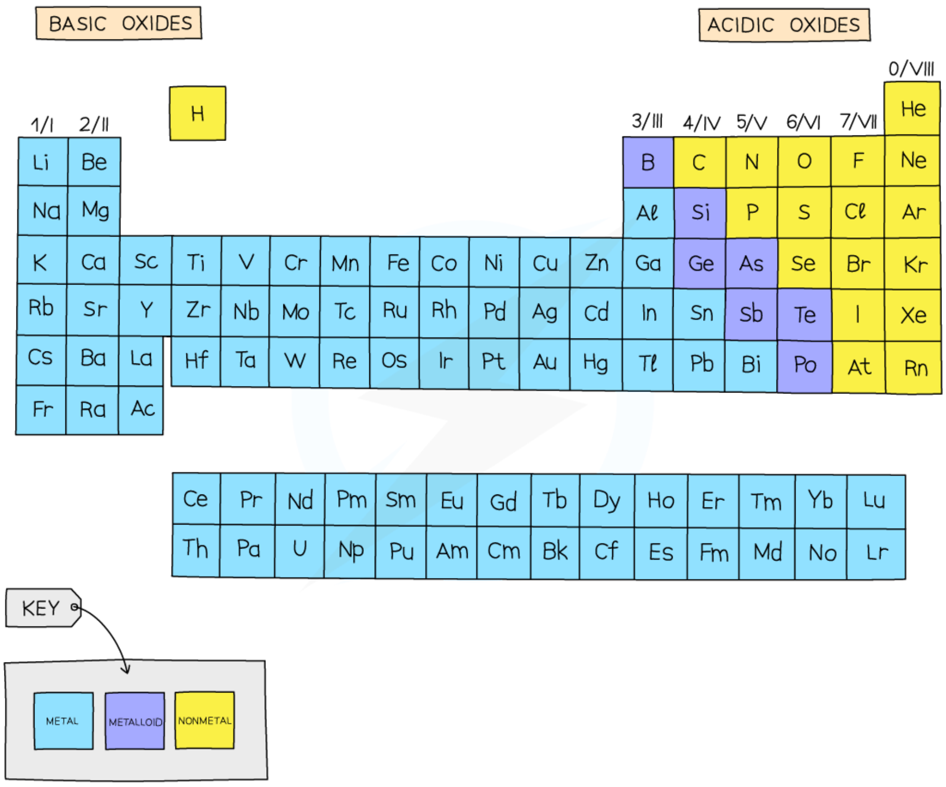
Acidic oxides

- Acidic oxides are formed when a **non-metal** element combines with oxygen
- They react with bases to form a salt and water
- When dissolved in water they produce an **acidic** solution with a **low** pH
- Common examples include CO_2, SO_2, NO_2 and SiO_2

Basic oxides

- Basic oxides are formed when a **metal** element combines with oxygen
- They react with acids to form a salt and water
- When dissolved in water they produce a **basic** solution with a **high** pH
- Common examples include CuO and CaO

Basic and acidic oxides



Metals form basic oxides while non-metals form acidic oxides

Amphoteric oxides

Extended tier only

Neutral oxides

- Some oxides do not react with either acids or bases and thus are said to be **neutral**
- Examples include N_2O, NO and CO

Amphoteric oxides

- Amphoteric oxides are a curious group of oxides that can behave as **both acidic** and **basic**, depending on whether the other reactant is an acid or a base
- In both cases salt and water are formed
- Two of the most common amphoteric oxides are **zinc oxide**, **ZnO** and **aluminium oxide**, **Al_2O_3**
- The **hydroxides** of both of these elements also behave amphoterically
- Example of aluminium oxide behaving as a base:

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$$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$$

- Example of aluminium oxide behaving as an acid:

$$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$$

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