IB Chemistry DP

YOUR NOTES

8. Acids & Bases

CONTENTS

8.1 Theories & Reactions of Acids & Bases
8.1.1 Brønsted–Lowry Acids & Bases
8.1.2 Conjugate Acid-Base Pairs
8.1.3 Characteristic Reactions of Acids
8.1.4 Neutralization
8.2 More About Acids
8.2.1 Acid-base Titrations
8.2.2 pH & [H^+]
8.2.3 Interpreting pH
8.2.4 The Ionic Product of Water
8.2.5 Acid-Base Calculations
8.2.6 pH Meters & Universal Indicator
8.2.7 Strong & Weak Acids & Bases
8.2.8 Comparing Strong & Weak Acids
8.3 Acid Deposition
8.3.1 Acid Deposition
8.3.2 Effects of Acid Deposition
8.3.3 Reducing Sulfur Oxide Emissions

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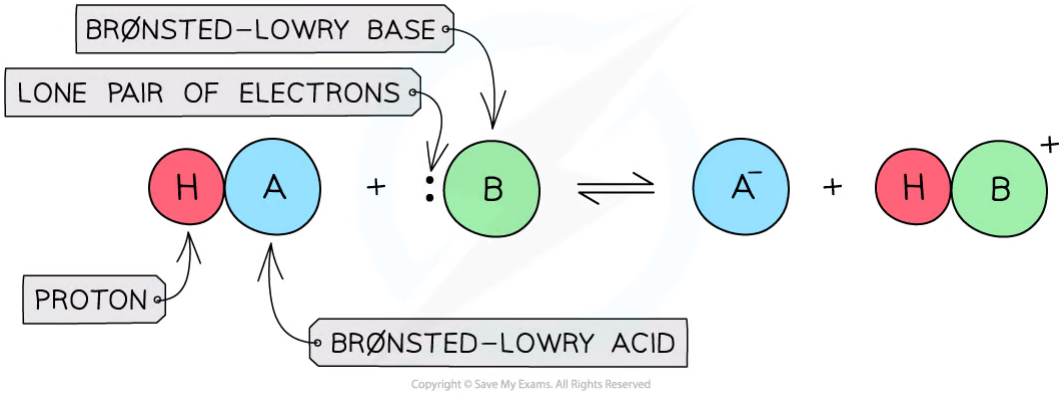
8.1 Theories & Reactions of Acids & Bases

8.1.1 Brønsted–Lowry Acids & Bases

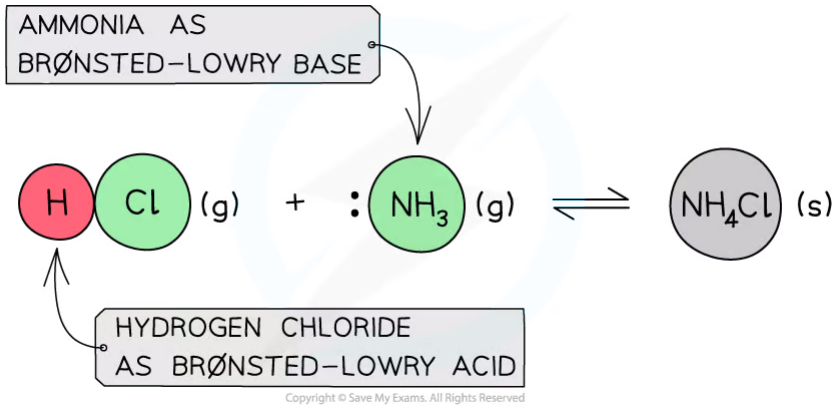
Brønsted–Lowry Acids & Bases

- **The Brønsted–Lowry Theory** defines acids and bases in terms of proton transfer between chemical compounds
- A **Brønsted–Lowry acid** is a species that **gives away** a **proton** (H^+)
- A **Brønsted–Lowry base** is a species that **accepts** a proton (H^+) using its **lone pair of electrons**



The diagram shows a Brønsted–Lowry acid which donates the proton to the Brønsted–Lowry base that accepts the proton using its lone pair of electrons

- The Brønsted-Lowry Theory is not limited to aqueous solutions only and can also be applied to reactions that occur in the gas phase



Example of a Brønsted–Lowry acid and base reaction in the gas state

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Worked Example

Identify the correct role of the species in the following reaction:

$$H_2PO_4^-(aq) + H_2O(l) \rightarrow HPO_4^{2-}(aq) + H_3O^+(aq)$$

	Brønsted–Lowry Acid	Brønsted–Lowry Base
A	$H_2PO_4^-$	H_2O
B	$H_2PO_4^{2-}$	$H_2PO_4^-$
C	$H_2PO_4^-$	H_3O^+
D	H_2O	$H_2PO_4^-$

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Answer:

The correct option is **A**.

- $H_2PO_4^-$ is donating a proton to H_2O, so $H_2PO_4^-$ must be an acid and H_2O must be a base

Exam Tip

An atom of hydrogen contains 1 **proton**, 1 electron and 0 neutrons. When hydrogen loses an electron to become **H^+** only a **proton** remains, which is why a H^+ ion is also called a proton.

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Conjugate Acid-Base Pairs

- A **Brønsted-Lowry acid** is a species that can donate a proton
- A **Brønsted-Lowry base** is a species that can accept a proton
- In a reaction equilibrium the products are formed at the same rate as the reactants are used

$$CH_3COOH\,(aq) + H_2O\,(l) \rightleftharpoons CH_3COO^-\,(aq) + H_3O^+\,(aq)$$

acid base conjugate base conjugate acid

- The reactant CH_3COOH is linked to the product CH_3COO^- by the transfer of a **proton** from the acid to the base
- Similarly, the H_2O molecule is linked to H_3O^+ ion by the transfer of a proton
- These pairs are therefore called **conjugate acid-base pairs**
- A **conjugate acid-base pair** is two species that are different from each other by a H^+ ion
  - **Conjugate** here means related
  - In other words, the acid and base are related to each other by one proton difference

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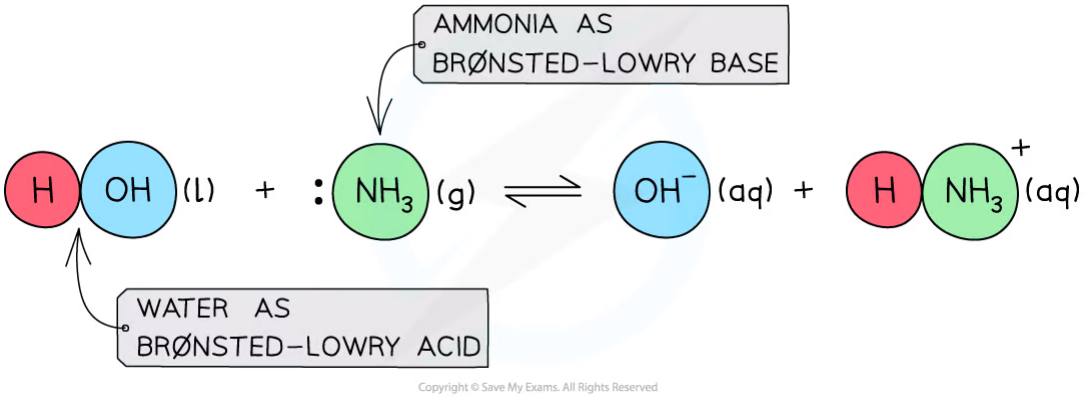
Amphiprotic Species

- Species that can act both as proton donors and acceptors are called **amphiprotic**
  - Eg. water as a Brønsted-Lowry acid



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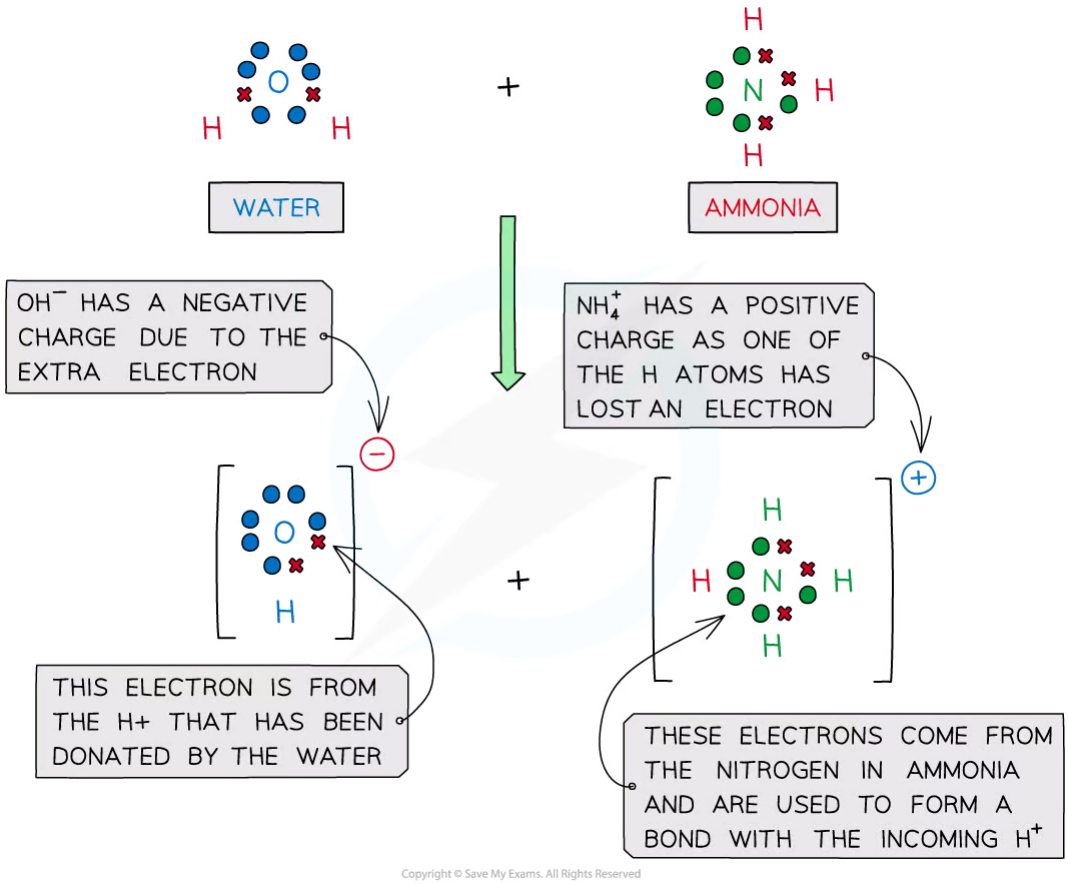
The diagram shows water acting as a Brønsted-Lowry acid by donating a proton to ammonia which accepts the proton using its lone pair of electrons



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The Lewis diagram for the reaction of water with ammonia to show how water acts as a Brønsted-Lowry acid and ammonia as a Brønsted-Lowry base

- Eg. water as a Brønsted-Lowry base

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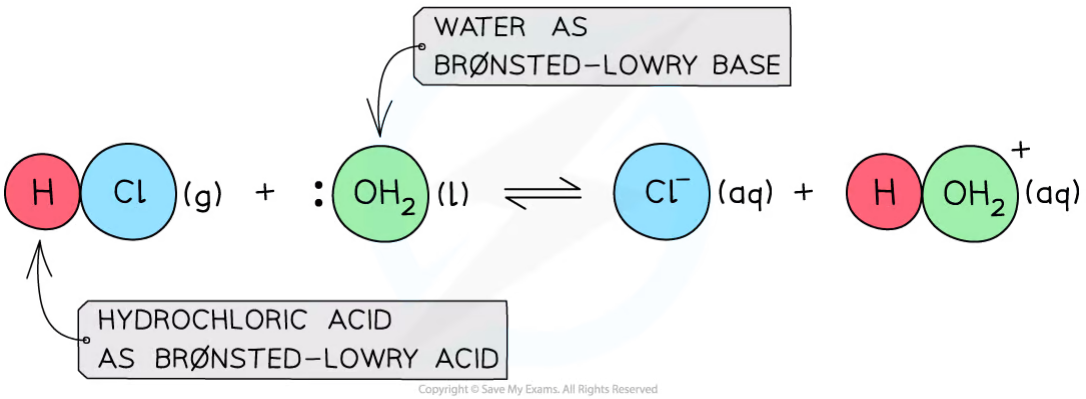
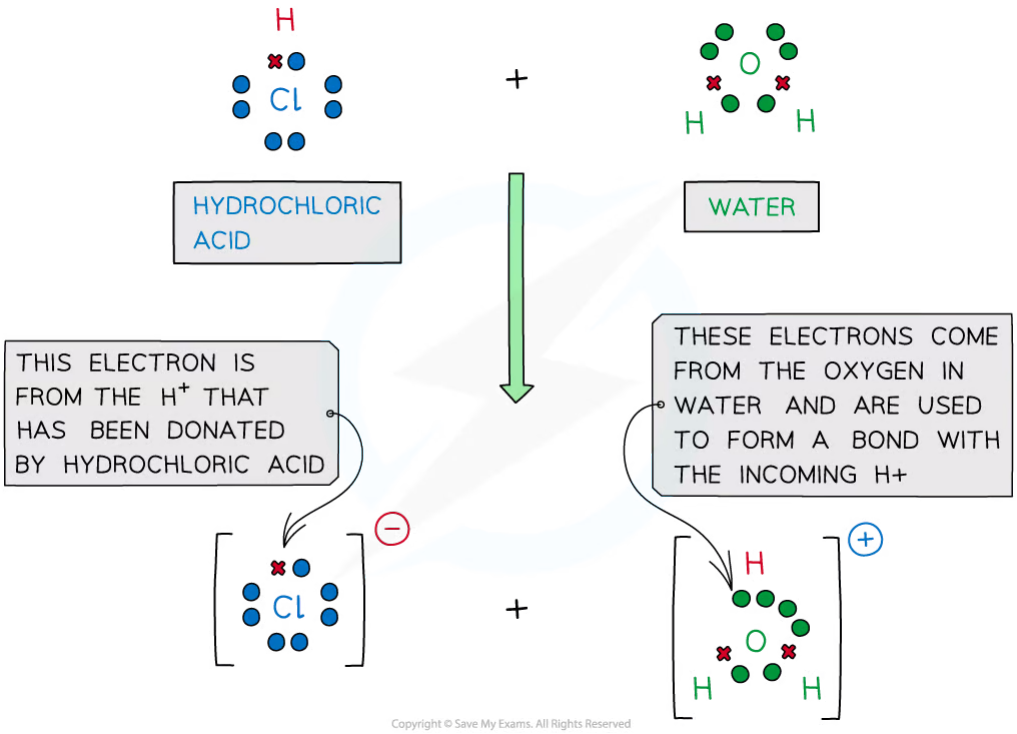


The diagram shows water acting as a Brønsted-Lowry base by accepting a proton from hydrochloric acid proton using its lone pair of electrons



The Lewis diagram for the reaction of water with hydrochloric acid to show how water acts as a Brønsted-Lowry base and ammonia as a Brønsted-Lowry acid

What is the difference between amphiprotic and amphoteric?

- A compound that is **amphoteric** means it has both basic and acidic character
  - When the compound reacts with an acid, it shows that it has basic character
  - When it reacts with a base, it shows that it's acidic
  - An example of this is aluminum oxide which reacts with both hydrochloric acid and sodium hydroxide:

$$Al_2O_3\,(s) + 6HCl\,(aq) \rightarrow 2AlCl_3\,(aq) + 3H_2O\,(l)$$

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$$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$$

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- When a compound is **amphiprotic**, it means it can act as a proton donor and as a proton acceptor
- Aluminium oxide is not amphiprotic, even though it is amphoteric

Amphiprotic versus Amphoteric Table

Amphiprotic	Amphoteric
The term amphiprotic describes a substance that can both accept and donate a proton or H^+	The term amphoteric refers to the ability to act as both an acid and a base
Amphiprotic substances can both accept or donate protons	Amphoteric substances can act as both an acid and a base
All amphiprotic substances are amphoteric	Not all amphoteric substances are amphiprotic

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Worked Example

In the equilibrium reaction shown below, which species are a conjugate acid-base pair? $CH_3CH_2CH_2COOH(aq) + H_2O(l) \rightleftharpoons CH_3CH_2CH_2COO^-(aq) + H_3O^+(aq)$

A. $CH_3CH_2CH_2COOH$ and H_2O

B. H_2O and H_3O^+

C. H_2O and $CH_3CH_2CH_2COO^-$

D. $CH_3CH_2CH_2COO^-$ and H_3O^+

Answer

The correct option in **B**

- A conjugate acid-base pair differ only by an H^+ ion

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Characteristic Reactions of Acids

Metals and acids

- The typical reaction of a metal and an acid can be summarized as

acid + metal → salt + hydrogen

- For example:

$2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$

hydrochloric acid + zinc → zinc chloride + hydrogen

- Clearly, the extent of reaction depends on the **reactivity** of the metal and the **strength** of the acid
- Very reactive metals would react dangerously with acids and these reactions are not usually carried out
- Metals low in **reactivity** do not react at all, for instance copper does not react with dilute acids
- **Stronger acids** will react **more vigorously** with metals than weak acids. What signs of reaction would be expected to be different between the two?
  - Faster reaction, seen as
    - more effervescence
    - the metal dissolves faster
  - More exothermic

Metals and oxides

- The reaction of an acid with a metal oxide forms two products:

acid + metal oxide → salt + water

- For example:

$2HCl(aq) + CaO(s) \rightarrow CaCl_2(aq) + H_2O(l)$

hydrochloric acid + calcium oxide → calcium chloride + water

Metals and hydroxides

- The reaction with a metal hydroxide and an acid follows the same pattern as an oxide:

acid + metal hydroxide → salt + water

- A suitable example might be:

$H_2SO_4(aq) + Mg(OH)_2(s) \rightarrow MgSO_4(aq) + 2H_2O(l)$

sulfuric acid + magnesium hydroxide → magnesium sulfate + water

Metals and carbonates

- The reaction between a metal carbonate and an acid produces three products:

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acid + metal carbonate → salt + water + carbon dioxide

- For example:

$$2HNO_3(aq) + CuCO_3(s) \rightarrow Cu(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

nitric acid + copper carbonate → copper nitrate + water + carbon dioxide

Metals and hydrogencarbonates

- The reaction between a metal hydrogencarbonate and an acid is the same as the carbonate reaction with a slight difference in stoichiometry:

acid + metal hydrogencarbonate → salt + water + carbon dioxide

- An example of this would be:

$$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

hydrochloric acid + sodium hydrogencarbonate → sodium chloride + water + carbon dioxide

Exam Tip

Make sure you learn the formulae of the common acids and bases and that you can write examples of balanced equations of their characteristic reactions

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Making Salts

- The acids and bases needed to make different salts can be deduced using the principles covered in the previous section
- The table below summarises these reactions

Making Salts Table

Type of salt	Ion	Acid needed	Formula	Base needed
Sulfates	SO_4^{2-}	sulfuric	H_2SO_4	metal oxide, hydroxide, carbonate or hydrogencarbonate
Nitrates	NO_3^-	nitric	HNO_3	
Chlorides	Cl^-	hydrochloric	HCl	
Ethanoates	CH_3COO^-	ethanoic	CH_3COOH	
Ammonium	NH_4^+	any	–	aqueous ammonia

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- Note that although some metals can be used to make salts, they are not classified as bases as water is not a product of the reaction

Worked Example

Which are the products of the reaction between zinc oxide and hydrochloric acid?

A. zinc chloride and carbon dioxide

B. zinc chloride, hydrogen gas and water

C. zinc, hydrogen gas and water

D. zinc chloride and water

Answer:

The correct option is **D**.

- Metal oxides when reacting with acids produce a salt and water as the only products

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Neutralization

- A neutralisation reaction is one in which an acid (pH <7) and a base/alkali (pH >7) react together to form water (pH = 7) and a salt

$$\text{ACID} + \text{BASE (ALKALI)} \longrightarrow \text{SALT} + \text{WATER}$$

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- The proton of the acid reacts with the hydroxide of the base to form water

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

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- The spectator ions which are not involved in the formation of water, form the salt

MAIN NEUTRALISATION REACTION:

$$HCl + NaOH \longrightarrow HOH + NaCl$$

ACID: HCl; BASE: NaOH; H_2O: HOH; SALT: NaCl

THE TWO INDIVIDUAL REACTIONS TAKING PLACE ARE:

1. $H^+ + OH^- \longrightarrow H_2O$
2. $Na^+ + Cl^- \longrightarrow NaCl$

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The diagram shows a neutralisation reaction of HCl and NaOH and the two individual reactions that take place to form the water and salt

- The name of the salt produced can be predicted from the acid that has reacted

Acid Reacted & Salt Table

Acid Reacted	Salt produced
Hydrochloric Acid	A Chloride
Nitric Acid	A Nitrate
Sulfuric Acid	A Sulfate

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Exam Tip

The enthalpy of neutralisation is the enthalpy change that occurs when an acid reacts with a base to form one mole of water. Since the reaction between strong acids and strong bases is the same regardless of the acid or base, it should be no surprise the enthalpy change is the same and is approximately -57 kJ mol^{-1}

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8.2 More About Acids

8.2.1 Acid-base Titrations

Acid-Base Titrations

- The steps involved in performing a **titration** and **titration calculation** are outlined in **Topic 1.2.9 Titrations**
- **Acid-base titrations** follow the same steps and are used to find the unknown concentrations of solutions of acids and bases
- **Acid-base indicators** give information about the change in chemical environment
- They change colour reversibly depending on the concentration of H^+ ions in the solution
- **Indicators** are weak acids and bases where the conjugate bases and acids have a different colour
- Many **acid-base indicators** are derived from plants, such as litmus

Common Indicators Table

Indicator	Colour in acid	Colour in alkali
Litmus	pink	blue
Methyl orange	red	yellow
Phenolphthalein	colourless	pink

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- A good indicator gives a very sharp colour change at the **equivalence point**
- In **titrations** is it not always possible to use two colour indicators because of this limitation, so for example litmus cannot be used successfully in a **titration**
- When **phenolphthalein** is used, it is usually better to have the base in the burette because it is easier to see the sudden and permanent appearance of a colour (pink in this case) than the change from a coloured solution to a colourless one

Exam Tip

Make sure you learn the colours of the common acid-base indicators

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pH & [H+]

- The acidity of an aqueous solution depends on the number of **H^+** (**H_3O^+**) ions in solution
- The **pH** is defined as:

$$pH = -\log_{10}[H^+]$$

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  - where $[H^+]$ is the concentration of H^+ in mol dm^{-3}

- The pH scale is a logarithmic scale with base 10
- This means that each value is 10 times the value below it. For example, pH 5 is 10 times more acidic than pH 6.
- pH values are usually given to 2 decimal places
- The relationship between concentration is easily seen on the following table:

pH & [H+] Table

$[H^+]$	Scientific notation	pH
1.0	10^0	0
0.1	10^{-1}	1
0.01	10^{-2}	2
0.001	10^{-3}	3
0.0001	10^{-4}	4
–	10^{-x}	x

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Worked Example

10.0 cm^3 of an aqueous solution of nitric acid of pH = 1.0 is mixed with 990.0 cm^3 of distilled water. What is the pH of the final solution?

A. 1

B. 2

C. 3

D. 10

Answer:

The correct option is **C**.

- The total volume after dilution is 1000.0 cm^3 so the concentration of H^+ has been **reduced** by a factor of 100 or 10^{-2}, which means an increase of 2 pH units
- The final solution is therefore **pH 3**

Exam Tip

Make sure you know how to use the antilog (base 10) feature on your calculator. On most calculators it is the 10^x button, but on other models it could be LOG^{-1}, ALOG or even a two-button sequence such as INV + LOG

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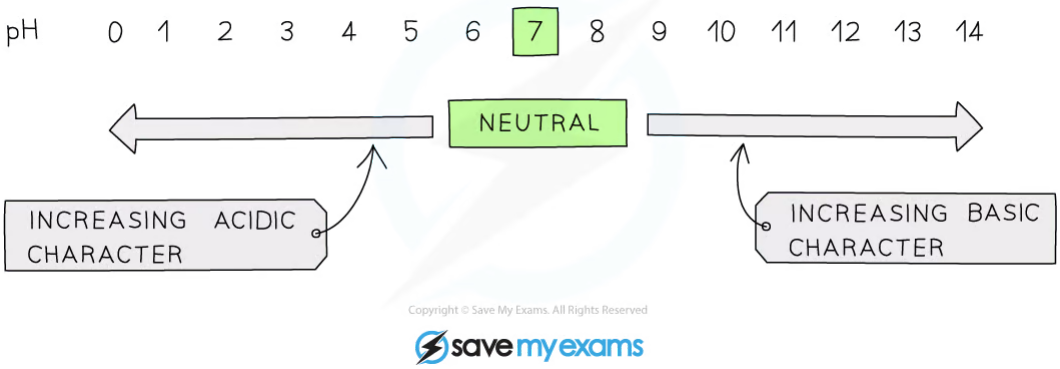
Interpreting pH

- The pH scale is a numerical scale that shows how **acidic** or **alkaline** a solution is
- The values on the pH scale go from 0–14 (extremely acidic substances have values of below 0)
- All acids have pH values of **below** 7, all alkalis have pH values **above** 7
- The **lower** the pH then the **more acidic** the solution is
- The **higher** the pH then the **more alkaline** the solution is



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The pH scale showing acidity, neutrality and alkalinity

pH of acids

- **Acidic** solutions (strong or weak) **always** have more H^+ than OH^- ions
- Since the concentration of **H^+** is always **greater** than the concentration of **OH^-** ions, $[H^+]$ is always **greater** than 10^{-7} mol dm^{-3}
- Using the pH formula, this means that the **pH of acidic solutions** is always **below** 7
- The higher the $[H^+]$ of the acid, the lower the pH

pH of bases

- **Basic** solutions (strong or weak) **always** have more OH^- than H^+ ions
- Since the concentration of **OH^-** is always **greater** than the concentration of **H^+** ions, $[H^+]$ is always **smaller** than 10^{-7} mol dm^{-3}
- Using the pH formula, this means that the **pH of basic solutions** is always **above** 7
- The higher the $[OH^-]$ of the base, the higher the pH

The pH of water

- Water at 298K has **equal amounts** of **OH^-** and **H^+ ions** with concentrations of **10^{-7} mol dm^{-3}**
- To calculate the pH of water, the following formula should be used:

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$$pH = -\log [H^+(aq)]$$

$[H^+(aq)]$ = CONCENTRATION OF H^+/H_3O^+ IONS

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$$pH = -\log(10^{-7})$$

$$= 7$$

- Thus, water has a pH of 7 at 298 K

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The Ionic Product of Water

pH of water

- An equilibrium exists in water where few water molecules dissociate into proton and hydroxide ions

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

- The equilibrium constant for this reaction is:

$$K_c = \frac{[H^+]\,[OH^-]}{[H_2O]}$$

$$K_c \times [H_2O] = [H^+]\,[OH^-]$$

- Since the concentration the H^+ and OH^- ions is very small, the concentration of water is considered to be a constant, such that the expression can be rewritten as:

$$K_w = [H^+]\,[OH^-]$$

Where K_w **(ionic product of water)** = $K_c \times [H_2O]$

= $10^{-14}\,mol^2\,dm^{-6}$ at 298K

- The product of the two ion concentrations is always $10^{-14}\,mol^2\,dm^{-6}$
- This makes it straightforward to see the relationship between the two concentrations and the nature of the solution:

$[H^+]$ & $[OH^-]$ Table

$[H^+]$	$[OH^-]$	Type of solution
0.1	1×10^{-13}	acidic
1×10^{-3}	1×10^{-11}	acidic
1×10^{-5}	1×10^{-9}	acidic
1×10^{-7}	1×10^{-7}	neutral
1×10^{-9}	1×10^{-5}	alkaline
1×10^{-11}	1×10^{-3}	alkaline
1×10^{-13}	0.1	alkaline

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Worked Example

What is the pH of a solution of potassium hydroxide, KOH(aq) of concentration 1.0×10^{-3} mol dm^{-3} ? $K_w = 1.0 \times 10^{-14}$ mol^2 dm^{-6}

A. 3

B. 4

C. 10

D. 11

Answer:

The correct option is **D**.

- Since $K_w = [H^+]\,[OH^-]$, rearranging gives $[H^+] = K_w \div [OH^-]$
- The concentration of $[H^+]$ is $(1.0 \times 10^{-14}) \div (1.0 \times 10^{-3}) = 1.0 \times 10^{-11}$ mol dm^{-3}
- So the **pH = 11**

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Acid-Base Calculations

- Using the relationships between pH, $[H^+]$ and $[OH^-]$ a variety of problems can be solved

$$\mathbf{pH = -\log [H^+]} \quad \text{and} \quad \mathbf{K_w = [H^+]\,[OH^-]}$$

- Test your understanding on the following worked examples:

Worked Example

1. The pH of a solution of phosphoric acid changes from 3 to 5. Deduce how the hydrogen ion concentration changes
2. Water from a pond was analysed and found to have a hydrogen ion concentration of 2.6×10^{-5} mol dm^{-3}. Calculate the pH of the pond water.
3. Determine the pH of a solution made by dissolving 5.00 g of potassium hydroxide in 250 cm^3 of distilled water

Answers:

Answer 1: The initial pH of the phosphoric acid is 3 which corresponds to a hydrogen ion concentration of 1×10^{-3} mol dm^{-3} :

$$[H+] = 10^{-pH}$$

$$\mathbf{[H+] = 1 \times 10^{-3}\ mol\ dm^{-3}}$$

The final pH is 5, which corresponds to 1×10^{-5} mol dm^{-3}

Therefore, the solution has decreased in $[H^+]$ concentration by 10^2 or 100 times

Answer 2: The pond water has $[H^+] = 2.6 \times 10^{-5}$ mol dm^{-3}.

$$\mathbf{pH = -\log [H+] = -\log(2.6 \times 10^{-5}) = 4.58}$$

Answer 3: Potassium hydroxide (M = 56.10 g mol^{-1}) is a strong base so the concentration of $[OH^-]$ is the same as the concentration of the solution as it fully dissociates:

$$KOH\,(s) \rightarrow K^+(aq) + OH^-(aq)$$

The concentration of KOH is

$$\frac{\frac{5.00}{56.10} \times 1000}{250\ cm^3} = 0.357 \text{ mol } dm^{-3} = [OH^-]$$

Using $K_w = [H^+][OH^-]$, and then rearranging $[H^+] = K_w/[OH^-]$

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$$[H^+] = \frac{1 \times 10^{-14}\ mol^2 dm^{-6}}{0.357\ mol\ dm^{-3}} = 2.80 \times 10^{-14}\ \text{mol dm}^{-3}$$

$pH = -\log(2.80 \times 10^{-14}) = 13.55$

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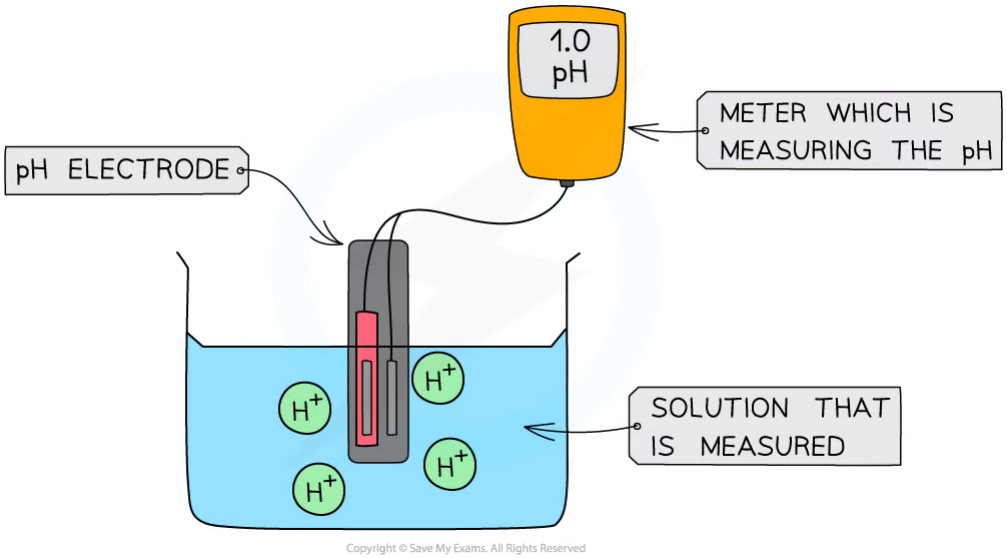
pH Meters & Universal Indicator

- The most **accurate** way to determine the pH is by reading it off a **pH meter**
- The pH meter is connected to the **pH electrode** which shows the pH value of the solution



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The diagram shows a digital pH meter that measures the pH of a solution using a pH electrode

- A less accurate method is to measure the pH using universal indicator paper
- The universal indicator paper is dipped into a solution of acid upon which the paper changes colour
- The colour is then compared to those on a chart which shows the colours corresponding to different pH values

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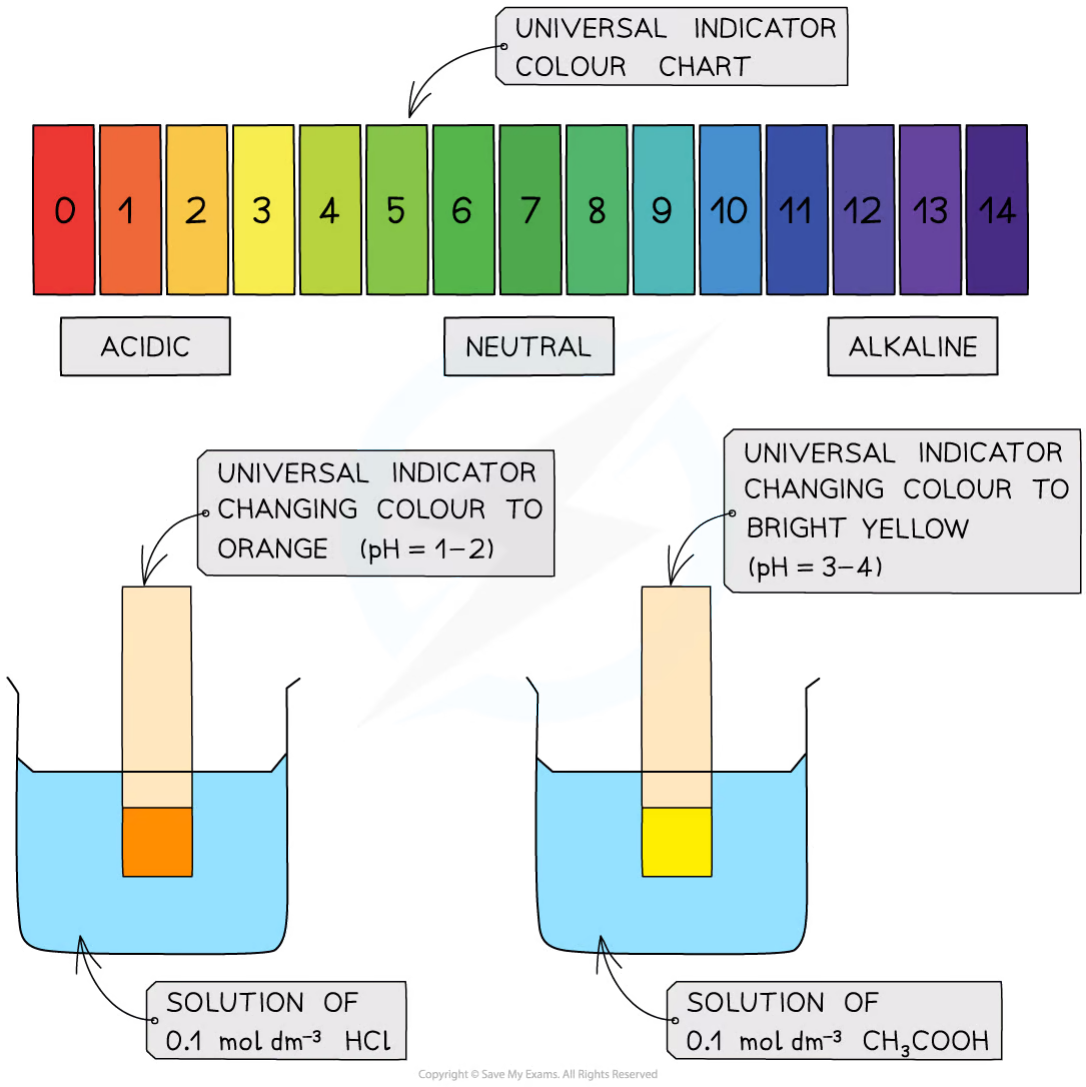

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The diagram shows the change in colour of the universal indicator paper when dipped in a strong (HCl) and weak (CH_3COOH) acid. The colour chart is used to read off the corresponding pH values which are between 1-2 for HCl and 3-4 for CH_3COOH

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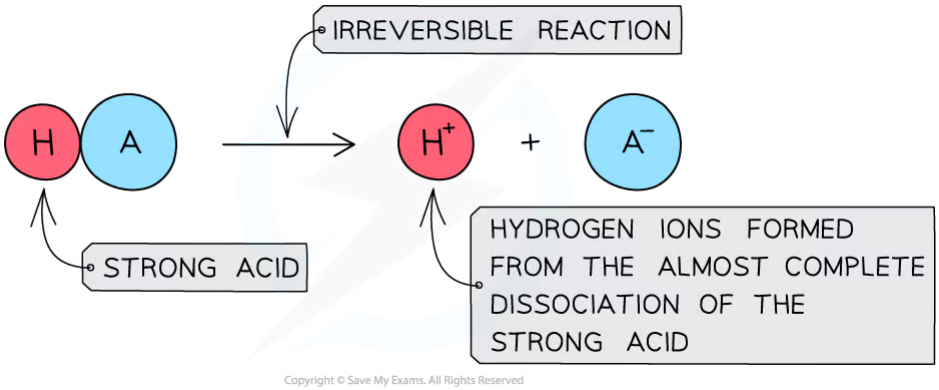
Strong & Weak Acids & Bases

Strong acids

- A **strong acid** is an acid that **dissociates** almost **completely** in aqueous solutions
  - HCl (hydrochloric acid), HNO_3 (nitric acid) and H_2SO_4 (sulfuric acid)
- The position of the equilibrium is so far over to the **right** that you can represent the reaction as an irreversible reaction



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The diagram shows the complete dissociation of a strong acid in aqueous solution

- The solution formed is **highly acidic** due to the high concentration of the H^+/H_3O^+ ions
- Since the **pH** depends on the concentration of H^+/H_3O^+ ions, the pH can be calculated if the concentration of the strong acid is known

$$pH = -\log [H^+(aq)]$$

$$[H^+(aq)] = \text{CONCENTRATION OF } H^+/H_3O^+ \text{ IONS}$$

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pH is the negative log of the concentration of H^+/H_3O^+ ions and can be calculated if the concentration of the strong acid is known using the stoichiometry of the reaction

Weak acids

- A **weak acid** is an acid that **partially** (or incompletely) **dissociates** in aqueous solutions
  - Eg. most organic acids (ethanoic acid), HCN (hydrocyanic acid), H_2S (hydrogen sulfide) and H_2CO_3 (carbonic acid)
- The position of the equilibrium is more over to the **left** and an equilibrium is established

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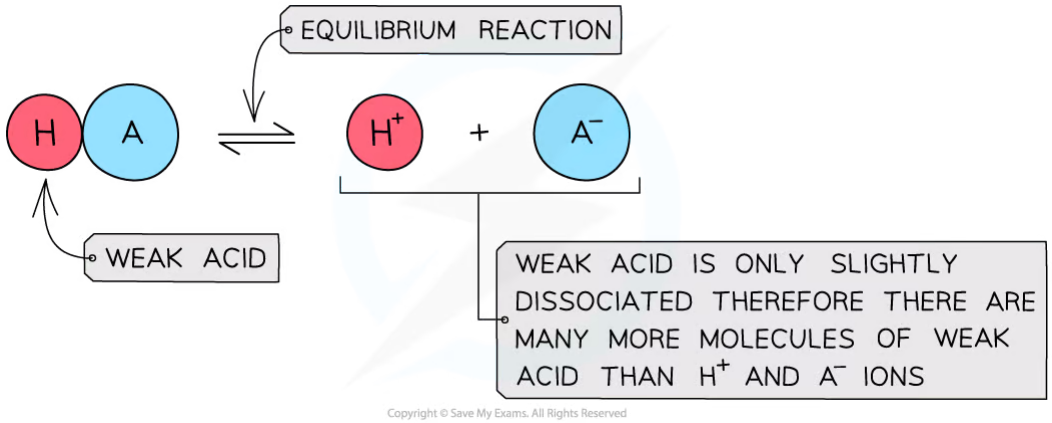

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The diagram shows the partial dissociation of a weak acid in aqueous solution

- The solution is **less acidic** due to the lower concentration of H^+/H_3O^+ ions
- Finding the pH of a weak acid requires using the acid dissociation constant, K_a but this not required at Standard Level, but only at Higher Level and is covered in Topic 18

Acid & Equilibrium Position Table

	Strong Acid	Weak Acid
Position of Equilibrium	Right	Left
Dissociation	Completely ($\longrightarrow$)	Partially ($\rightleftharpoons$)
H^+ concentration	High	Low
pH	Use [strong acid] for $[H^+]$	Use K_a to find $[H^+]$
Examples	HCl HNO_3 H_2SO_4 (first ionisation)	Organic acids (ethanoic acid) HCN H_2S H_2CO_3

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Strong bases

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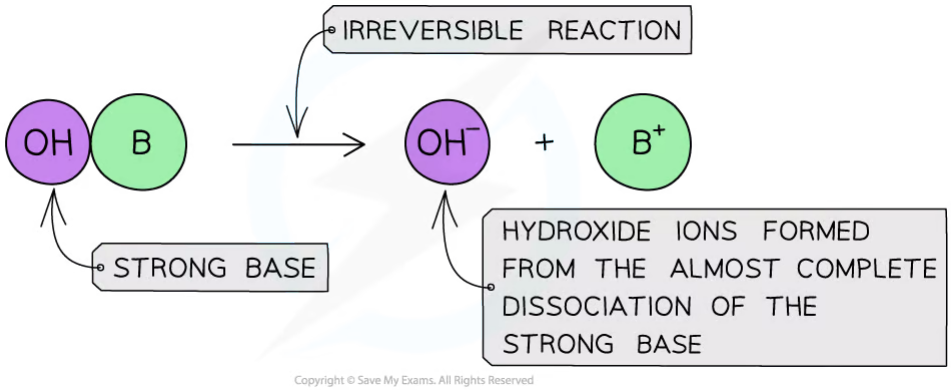
- A **strong base** is a base that dissociates almost completely in aqueous solutions
  E.g. group 1 metal hydroxides such as NaOH (sodium hydroxide)
- The position of the equilibrium is so far over to the right that you can represent the reaction as an irreversible reaction

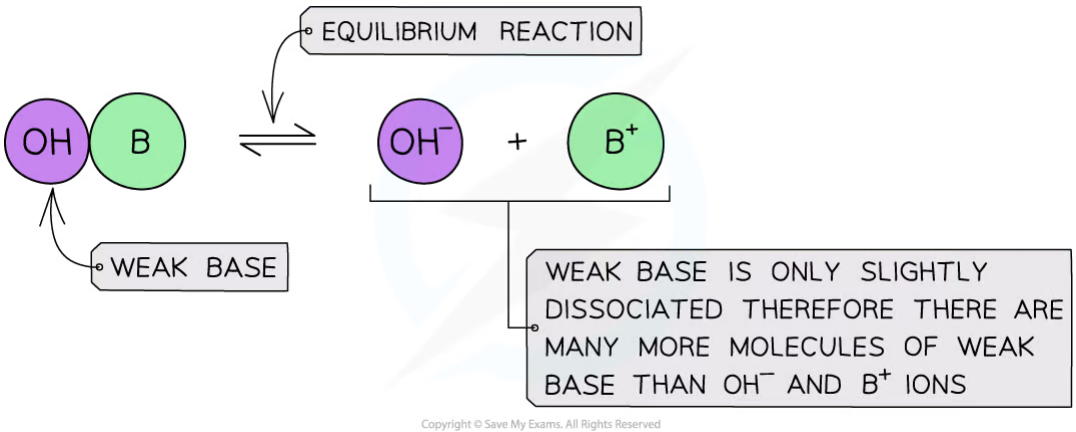


The diagram shows the complete dissociation of a strong base in aqueous solution

- The solution formed is highly basic due to the high concentration of the OH^- ions

Weak bases

- A **weak base** is a base that **partially** (or incompletely) **dissociates** in aqueous solutions
  - NH_3 (ammonia), amines and some hydroxides of transition metals
- The position of the equilibrium is more to the **left** and an equilibrium is established



The diagram shows the partial dissociation of a weak base in aqueous solution

- The solution is **less basic** due to the lower concentration of OH^- ions

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Base & Equilibrium Position Table

	Strong Base	Weak Base
Position of Equilibrium	Right	Left
Dissociation	Completely ($\longrightarrow$)	Partially ($\rightleftharpoons$)
OH^- concentration	High	Low
Examples	Group 1 metal hydroxides	NH_3 Amines Some transition metal hydroxides

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Conjugate Pairs & Acid-Base Strength

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- The conjugate base of HCl is the chloride ion, Cl^-, but since the reverse reaction is virtually non-existent the chloride ion must be a very weak conjugate base

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

acid **conjugate base**

- In general **strong acids** produce **weak conjugate bases** and **weak acids** produce **strong conjugate bases**
- A strong base is also fully ionized and is a good proton acceptor
- For example the hydroxide ion is a strong base and readily accepts protons:

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

- The conjugate acid of the hydroxide ion is water, which is a weak conjugate acid
- In general **strong bases** produce **weak conjugate acids**

Exam Tip

Hydrogen ions in aqueous solutions can be written as either as H_3O^+ or as H^+ however, if H_3O^+ is used, H_2O should be included in the chemical equation: **$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$** OR **$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$** Some acids contain two replaceable protons (called '**dibasic**') – for example, H_2SO_4 (sulfuric acid) has two ionisations: H_2SO_4 acts as a strong acid: **$H_2SO_4 \rightarrow H^+ + SO_4^-$**$HSO_4^-$ acts as a weak acid: **$HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$**The second ionisation is only partial which is why the concentration of 1 mol dm^{-3} sulfuric acid is not 2 mol dm^{-3} in H^+ ions Also, don't forget that the terms **strong** and **weak** acids and bases are related to the **degree of dissociation** and not the **concentration.**The appropriate terms to use when describing **concentration** are **dilute** and **concentrated.**

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Comparing Strong & Weak Acids

- Strong and weak acids can be distinguished from each other by their:
  - **pH value** (using a pH meter or universal indicator)
  - **Electrical conductivity**
  - **Reactivity**

pH value

- An acid **dissociates** into H^+ in solution according to:

$$HA \rightarrow H^+ + A^-$$

- The **stronger** the acid, the **greater** the **concentration of H^+** and therefore the **lower the pH**

pH value of a Strong Acid & Weak Acid Table

Acid	pH of 0.1 mol dm^{-3} solution
HCl (strong)	1
CH_3COOH (weak)	2.9

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Electrical conductivity

- Since a **stronger acid** has a **higher concentration of H^+** it **conducts electricity** better
- Stronger acids therefore have a greater **electrical conductivity**
- The electrical conductivity can be determined by using a **conductivity meter**
- Like the pH meter, the conductivity meter is connected to an electrode
- The conductivity of the solution can be read off the meter

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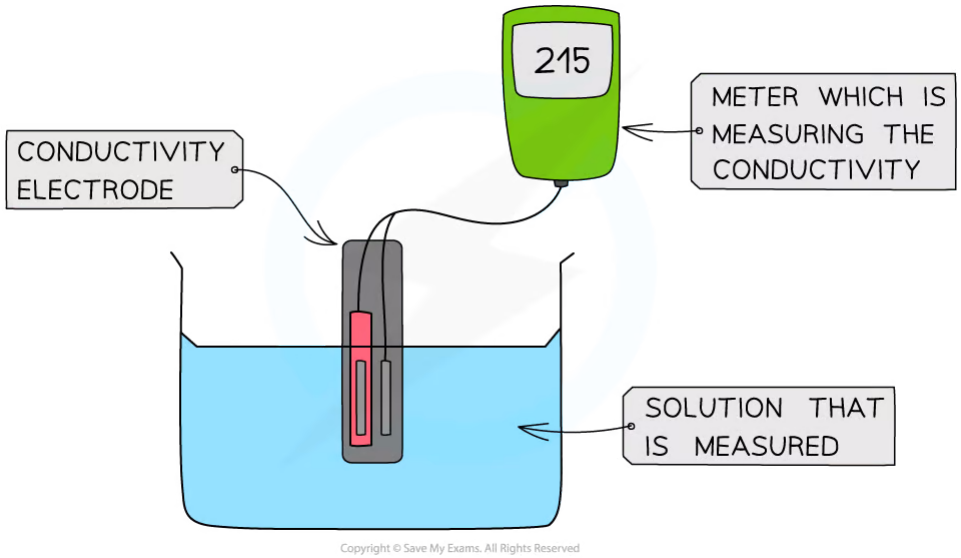

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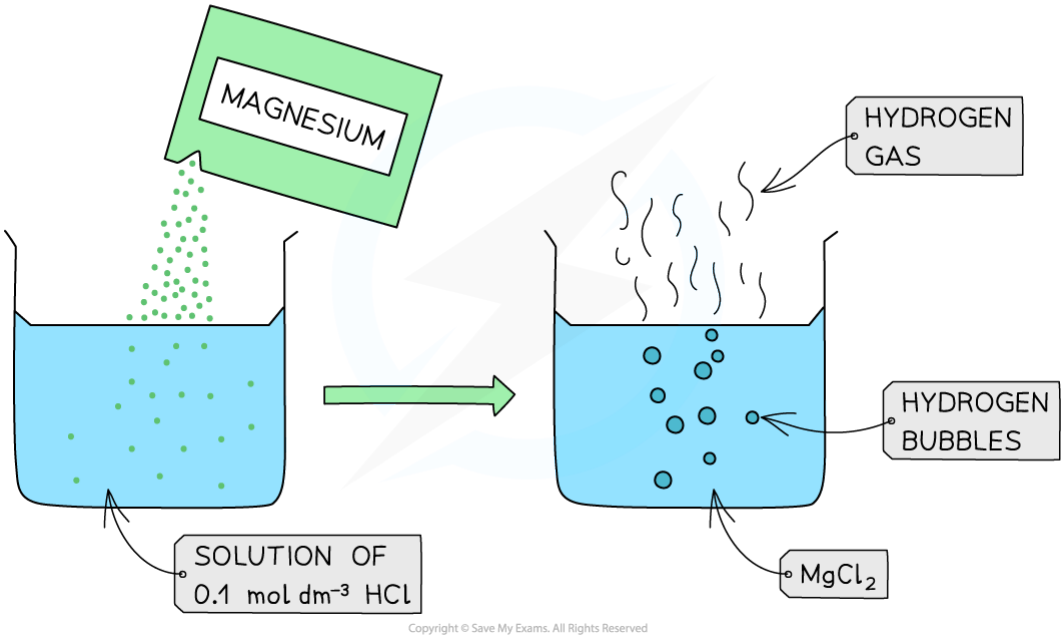
The diagram shows a digital conductivity meter that measures the electrical conductivity of a solution using an electrode

Reactivity

- Strong and weak acids of the **same concentrations** react differently with reactive metals
- This is because the concentration of H^+ is greater in strong acids compared to weak acids
- The greater H^+ concentration means that more H_2 gas is produced in a shorter time

$$Mg\,(s) + 2HCl\,(aq) \longrightarrow MgCl_2\,(aq) + H_2\,(g)$$



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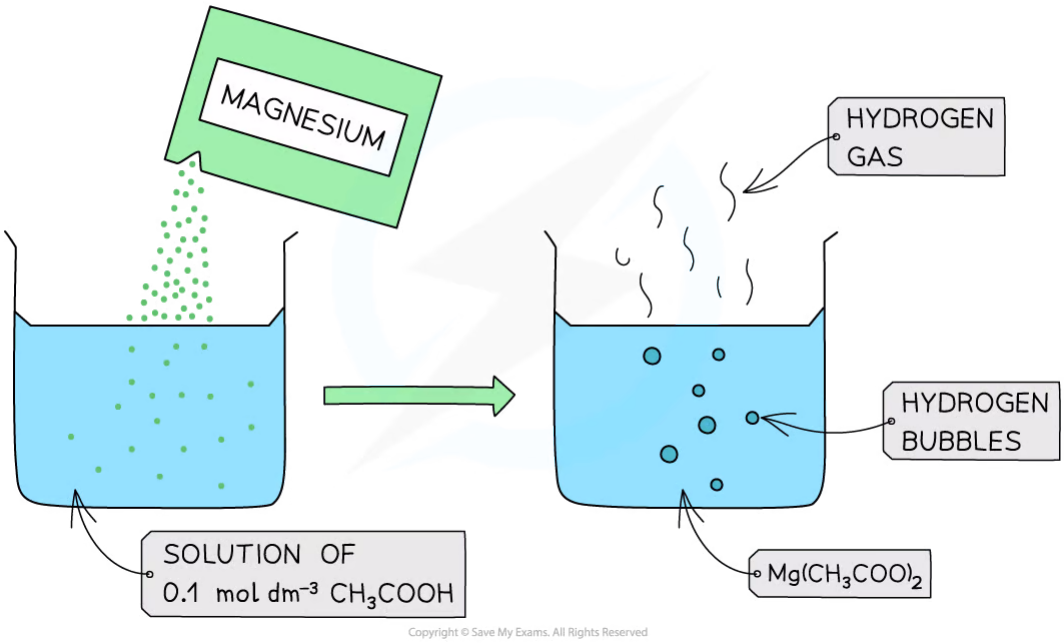
The diagram shows the reaction of 0.1 mol dm^{-3} of a strong acid (HCl) with Mg. The reaction produces a lot of bubbles and hydrogen gas due to the high concentration of H$^+$ present in solution

$$Mg\,(s) + 2CH_3COOH\,(aq) \longrightarrow Mg(CH_3COO)_2\,(aq) + H_2\,(g)$$



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The diagram shows the reaction of 0.1 mol dm^{-3} of a weak acid (CH_3COOH) with Mg. The reaction produces fewer bubbles of hydrogen gas due to the lower concentration of H$^+$ present in solution

- Similar observations would be made in the reaction between strong and weak acids with **carbonates** and **hydrogencarbonates**, although the gas given off this time is carbon dioxide
- With **oxides** and **hydroxides**, there may not be a lot of visible changes although it is likely that they would **dissolve faster** in a strong acid than in a weak acid
- These reactions are also likely to produce **larger enthalpy changes** which could be seen in **higher temperature rises**

Exam Tip

The above-mentioned properties of strong and weak acids depend on their ability to dissociate and form H$^+$ ions. Stronger acids dissociate more, producing a greater concentration of H$^+$ ions and therefore showing lower pH values, greater electrical conductivity and more vigorous reactions with reactive metals.

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8.3 Acid Deposition

8.3.1 Acid Deposition

Acid Deposition

What is acid deposition?

- Rain is naturally acidic because of dissolved CO_2 which forms **carbonic acid**

$$H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq)$$

- **Carbonic acid** is a **weak acid** and dissociates in the following equilibrium reaction giving a **pH of 5.6**

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

- For that reason **acid rain** is defined as rain with a pH of **below** 5.6
- **Acid deposition** includes all processes by which acidic components leave the atmosphere
  - This could be gases or precipitates
- There are two types of deposition: **wet acid deposition** and **dry acid deposition**
  - **Wet acid deposition** refers to rain, snow, sleet, hail, fog, mist and dew
  - **Dry acid deposition** refers to acidic particles and gases that fall to the ground as dust and smoke
- **Acid deposition** is formed when nitrogen or sulfur oxides dissolve in water to form HNO_3, HNO_2, H_2SO_4 and H_2SO_3

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Acid Deposition Equations

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Formation of sulfur based acids

- **Fossil fuels** are often contaminated with small amounts of **sulfur impurities**
- When these contaminated fossil fuels are **combusted**, the sulfur in the fuels get **oxidised** to **sulfur dioxide**

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

- Sulfur dioxide may be further oxidised to sulfur trioxide

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

- The sulfur dioxide and sulfur trioxide then dissolve in rainwater droplets to form **sulfurous acid** and **sulfuric acid**

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

- These acids are components of acid rain which has several damaging impacts on the environment

Formation of acid rain by nitrogen oxides

- The temperature in an internal combustion engine can reach over 2000 °C
- Here, nitrogen and oxygen, which at normal temperatures don't react, combine to form **nitrogen monoxide**:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$

- **Nitrogen monoxide** reacts further forming **nitrogen dioxide**:

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$$

- **Nitrogen dioxide** gas reacts with rain water to form a mixture of **nitrous** and **nitric acids**, which contribute to **acid rain**:

$$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$

- Lightning strikes can also trigger the formation of nitrogen monoxide and nitrogen dioxides in air
- Nitrogen dioxide gas reacts with rain water and more oxygen to form **nitric acid**

$$4NO_2(g) + 2H_2O(l) + O_2(g) \rightarrow 4HNO_3(aq)$$

- When the clouds rise, the temperature decreases, and the droplets get larger
- When the droplet containing these acids are heavy enough, they will fall down as **acid rain**

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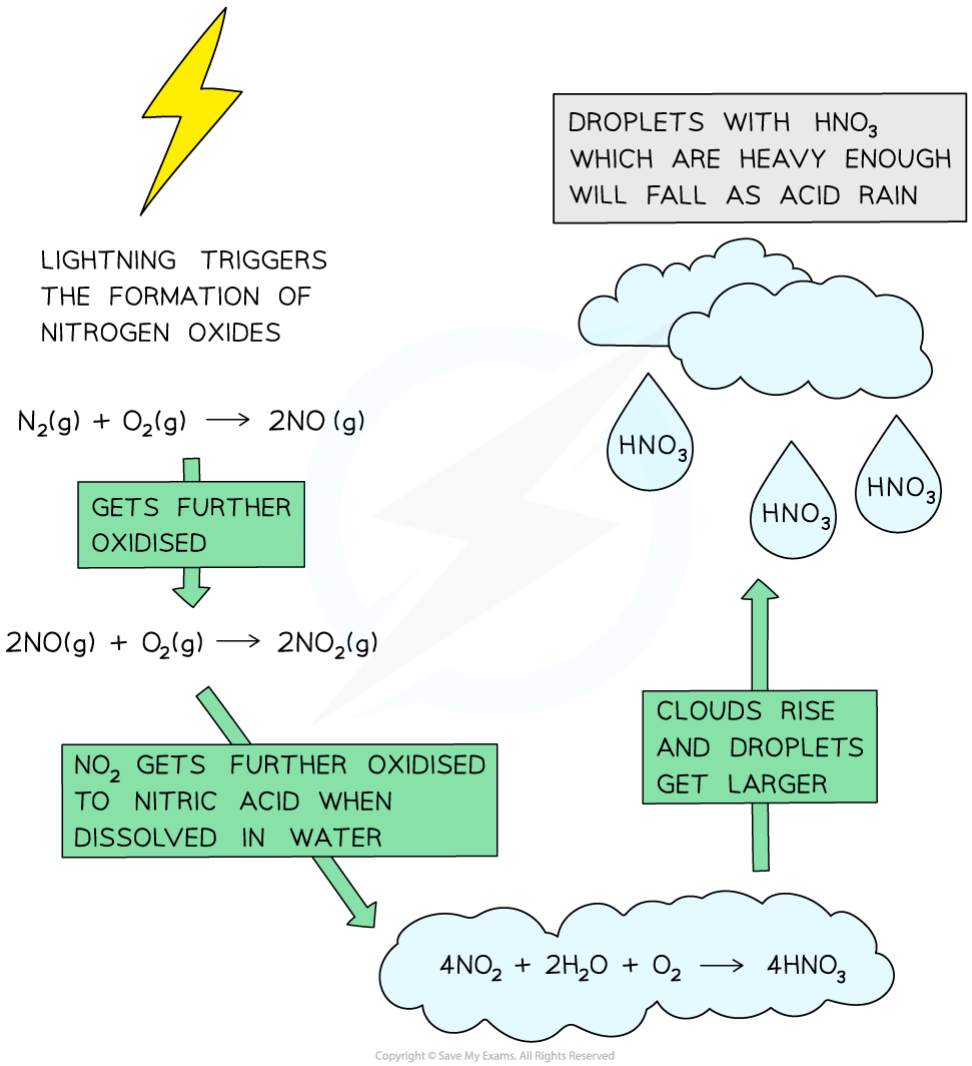

The diagram shows the formation of acid rain by the oxidation of nitrogen dioxide

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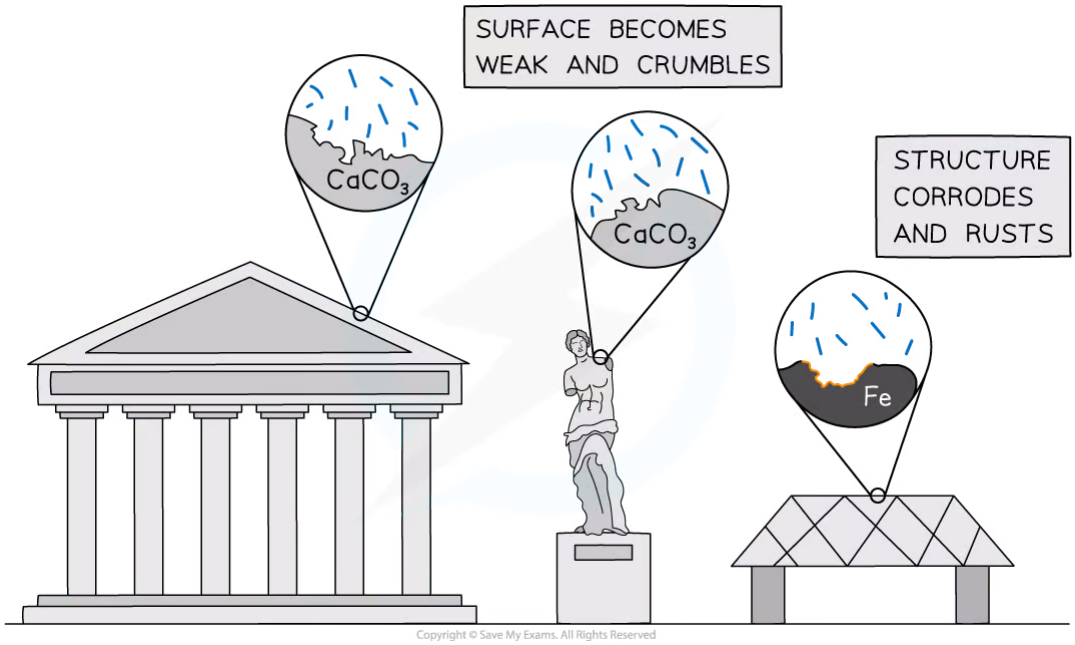
Effects of acid deposition

Effect on materials

- **Acid deposition** can react with metals and rocks (such as **limestone**) causing buildings and statues to get damaged



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The impact of acid deposition can be seen on buildings, statues and metallic structures, particularly in polluted cities

- Limestone and marble both contain calcium carbonate
- The calcium carbonate reacts with sulfuric or nitric acids causing stonework to corrode and weaken:

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + CO_2(g) + H_2O(l)$$

$$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

- In addition dry deposition also affects limestone:

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + CO_2(g) + H_2O(l)$$

- Metallic structures, such as iron bridges, are vulnerable to corrosion by wet and dry deposition forming the salt of the metal:

$$Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$$

$$Fe(s) + SO_2(g) + O_2(g) \rightarrow FeSO_4(s)$$

Effect on plants

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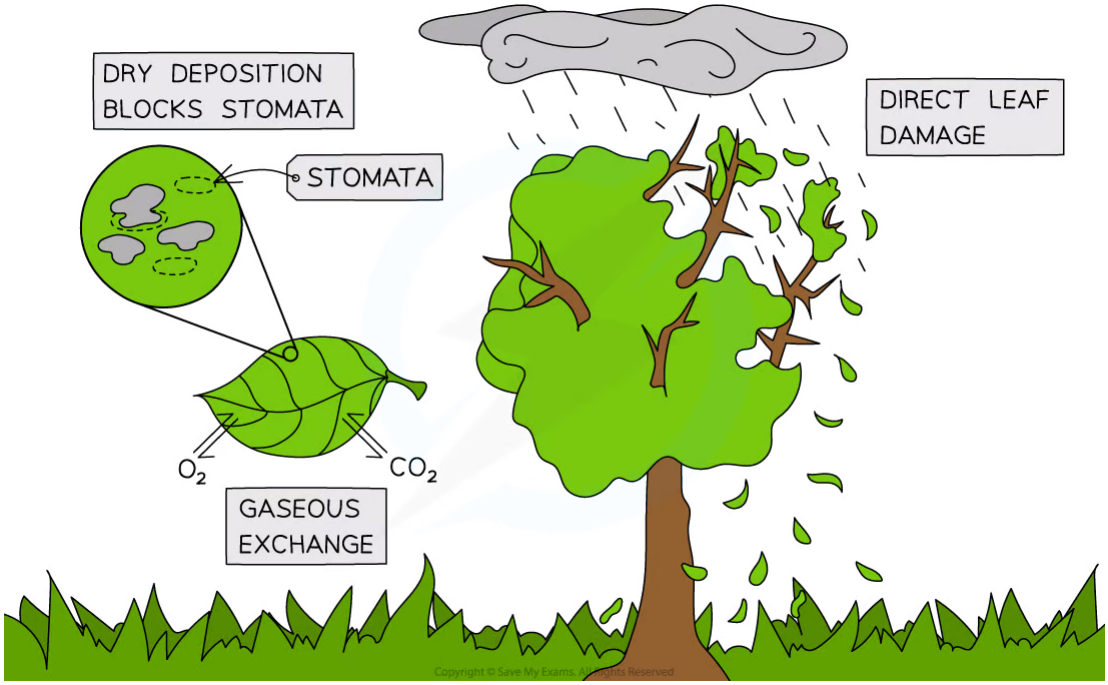
- Apart from **acid deposition** directly falling on leaves and killing plants, acid particulates can block stomata (plant pores) and prevent gaseous exchange



The impact of acid deposition on plant life

- Acid rain can fall on soils and release important minerals such as magnesium, calcium and potassium which are leached (washed out) from soils and are therefore unavailable to plants
- Aluminium ions released from rocks are toxic to many plants and damage their roots

Effect on water

- When **acid rain** falls on rivers and lakes the pH can fall to levels that are unable to support life

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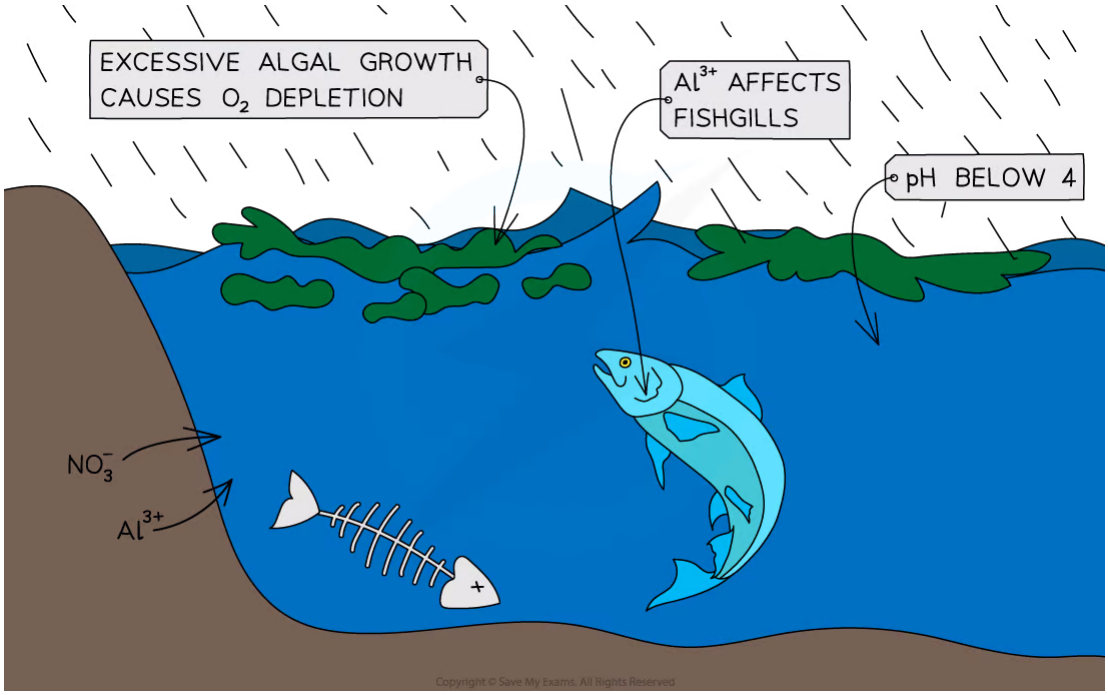

The impact of acid deposition on waterways

- Below about pH 4, aluminium ions are released from rocks when they are held as aluminium hydroxide

$$Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$$

- Alumnium ions are toxic to fish as they damage the gills and prevent fish from efficiently absorbing oxygen
- Nitrate ions from nitric acid in **acid rain** can contibute to over-fertilization of waterways and lead to **eutrophication**
  - **Eutrophication** is excessive algal growth that results in oxygen depletion and stagnation of waterways

Effect on human health

- Although acid rain is too dilute to cause any direct impact on the skin, acidic particulates in the air can increase the risk of respiratory diseases such as bronchitis, asthma and emphysema
- When acid rain comes into contact with metal pipes there is an increased risk that toxic metal ions will be released into the water supply such as Cu^{2+}, Al^{3+} and Pb^{2+}

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Reducing Sulfur Oxide Emissions

- The removal of sulfur from fossil fuels can either take place **pre-combustion** or **post-combustion**
- The oxides of sulfur, **SO_2** and **SO_3**, are both **acidic** and **toxic gases**
- **Sulfur dioxide** is produced naturally during volcanic eruptions, but large quantities have been and continue to be emitted by burning coal, oil and natural gas

Pre-combustion

- **Pre-combustion** of sulfur takes place for coal and petroleum, although it is expensive to remove all the sulfur, so a small percentage often remains
  - For example, the average sulfur content of gasoline is 347ppm (this is the same as 347 mg per litre)
- It is essential to remove most of the sulfur as it damages the workings of internal combustion engines
- The sulfur is removed by reacting it with hydrogen in a process called **hydrodesulfurization**
- The sulfur is recovered and used in the manufacture of sulfuric acid

Post-combustion

- Post-combustion is carried out on in coal-fired power stations
- The waste gases from burning the coal contain sulfur dioxide
- The waste gases are passed through a wet slurry of calcium oxide and calcium carbonate which react with the SO_2 and produce calcium sulfate

$$CaO(s) + SO_2(g) + \tfrac{1}{2}O_2(g) \rightarrow CaSO_4(s)$$

$$CaCO_3(s) + SO_2(g) + \tfrac{1}{2}O_2(g) \rightarrow CaSO_4(s) + CO_2(g)$$

- The calcium sulfate is also known as **gypsum** and is used to make plasterboard and other useful building materials

What does the future hold?

- Global policies working towards combating **global warming** will have the additional benefit in **reducing acid deposition**
- As we switch away from burning fossil fuels for energy there will be a **fall** in the emission of oxides of sulfur and nitrogen, which can only be a good thing for the environment
- Ultimately reducing the emission of primary pollutants is achieved by greater use of **renewable energy sources**, greater use of public transport and more **efficient energy transfer systems**

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