SL IB Chemistry

Proton Transfer Reactions

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Brønsted–Lowry Acids & Bases

Brønsted–Lowry Acids & Bases

What are Brønsted-Lowry acids and 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**
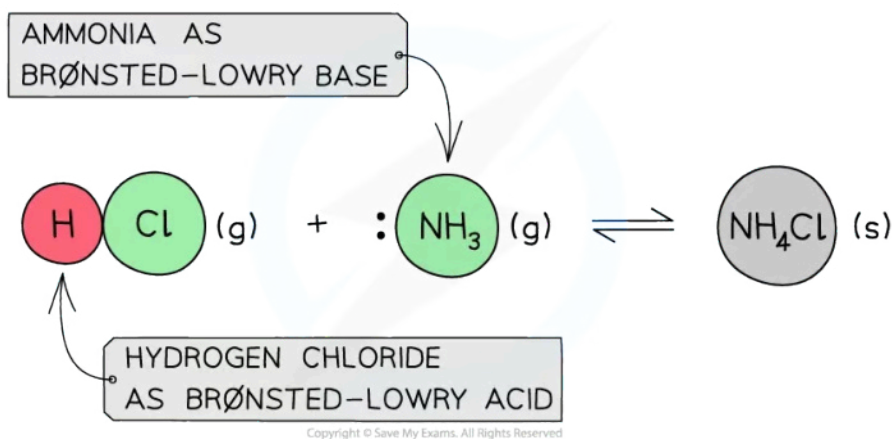
Equilibrium of a Brønsted–Lowry acid and base
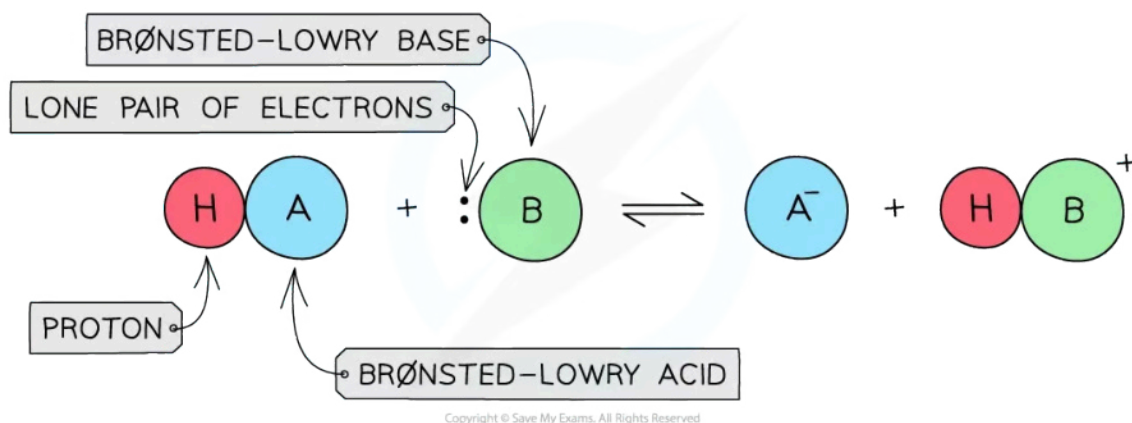
The diagram shows a Brønsted–Lowry acid which donates the proton to the Brønsted–Lowry base that accepts the proton using its lone pair of electrons

- The Brønsted-Lowry Theory is not limited to aqueous solutions only and can also be applied to reactions that occur in the gas phase
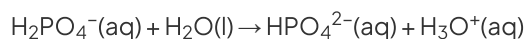
Diagram to show how hydrochloric acid acts as a Brønsted-Lowry acid and ammonia acts as a Brønsted–Lowry base

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

✏️ 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	$H_2PO_4^-$

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

> ### 💡 Examiner 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.

Your notes

Conjugate Acids & Bases

Conjugate Acids & Bases

- 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 an equilibrium reaction, the products are formed at the same rate as the reactants are used

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

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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✏️ **Worked example**

In the equilibrium reaction shown below, which species are a conjugate acid–base pair?

$$CH_3CH_2CH_2COOH\,(aq) + H_2O\,(l) \rightleftharpoons CH_3CH_2CH_2COO^-\,(aq) + H_3O^+\,(aq)$$
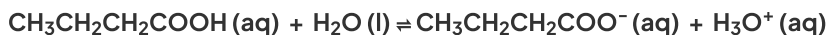
A. $CH_3CH_2CH_2COOH$ and H_2O

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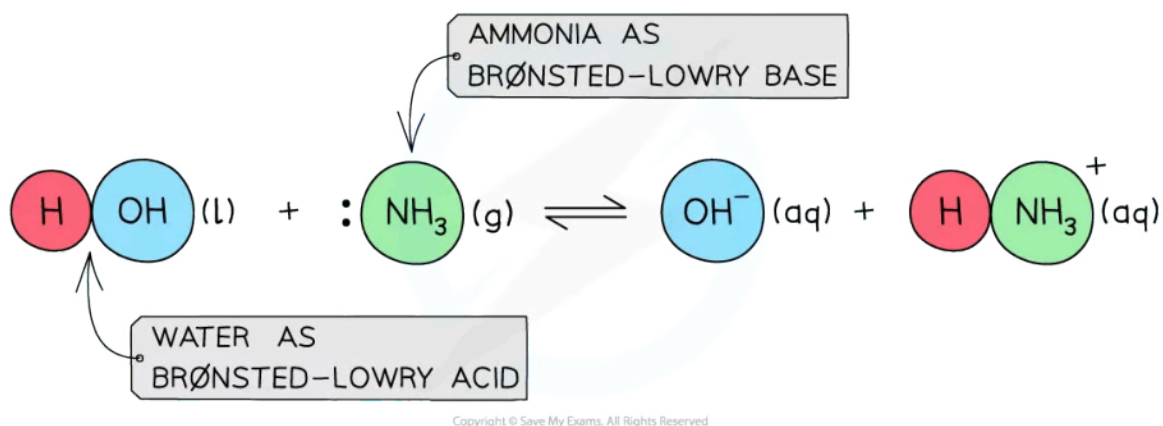   - A conjugate acid–base pair differ only by an H^+ ion

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Amphiprotic Species

Amphiprotic Species

- Species that can act both as proton donors and acceptors are called **amphiprotic**
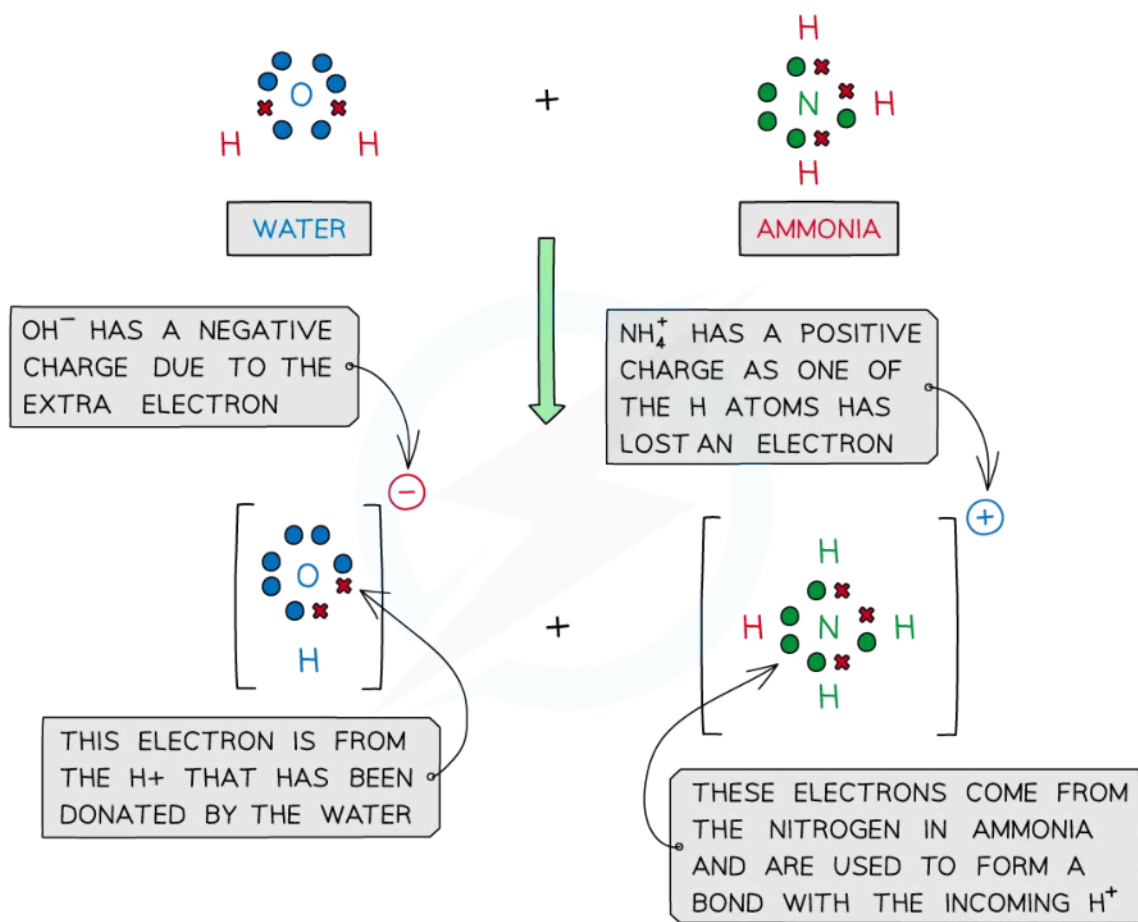
Water as a Brønsted–Lowry acid

Diagram to show how water is amphiprotic

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
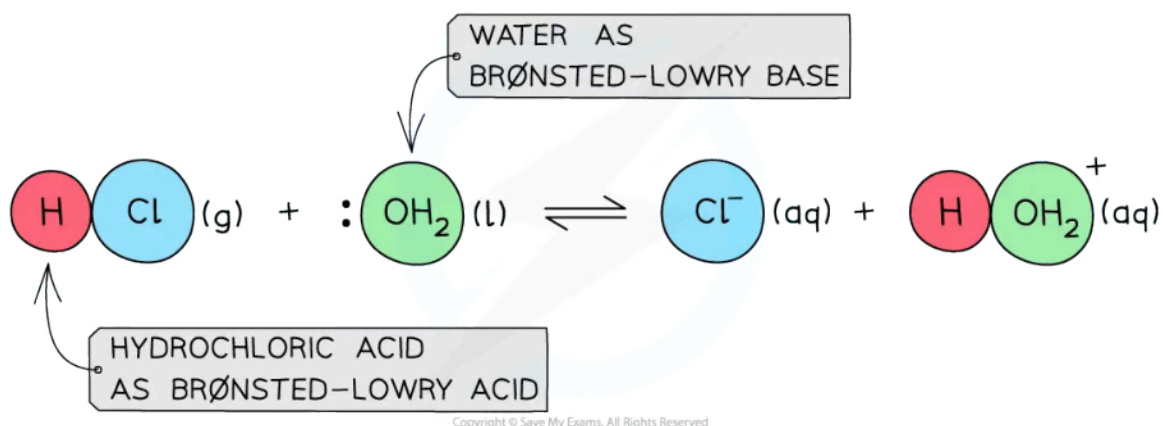
Lewis diagram for the reaction between water and ammonia

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
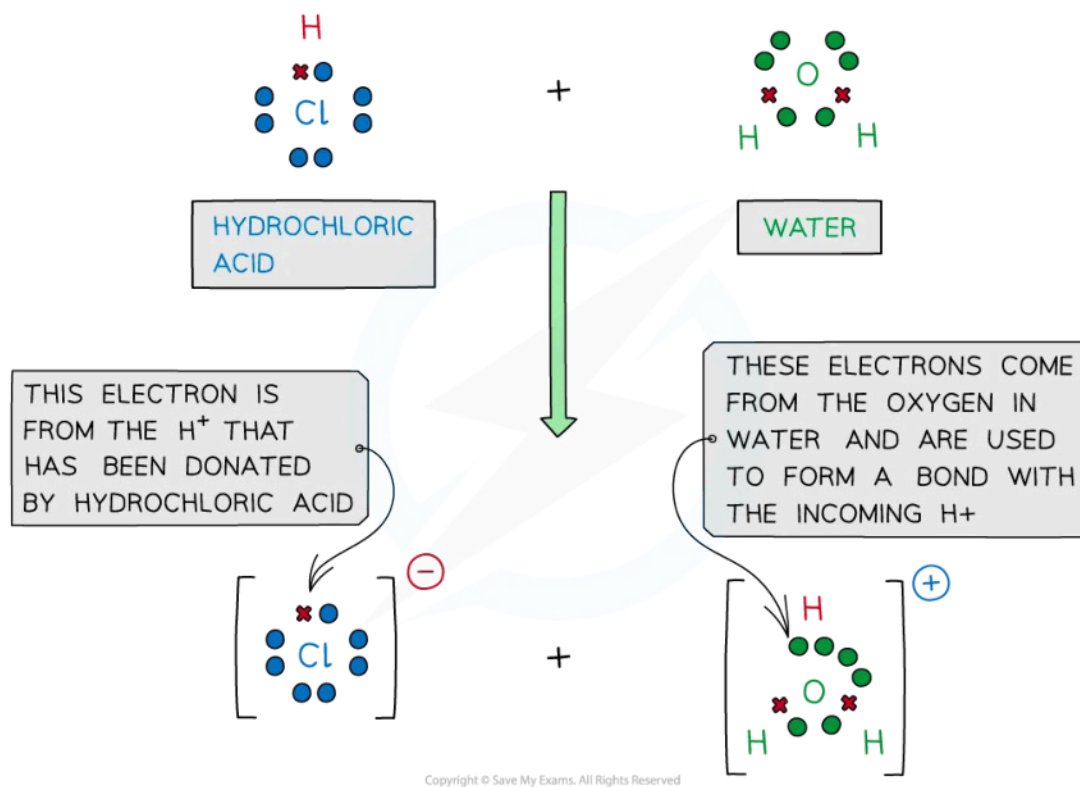
Water as a Brønsted–Lowry base

Diagram to show how water is amphiprotic

Your notes

WATER AS BRØNSTED–LOWRY BASE

HYDROCHLORIC ACID AS BRØNSTED–LOWRY ACID

The diagram shows water acting as a Brønsted–Lowry base by accepting a proton from hydrochloric acid proton using its lone pair of electrons
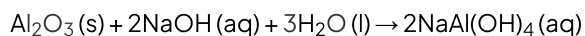
Lewis diagram for the reaction between water and hydrochloric acid

HYDROCHLORIC ACID

WATER

THIS ELECTRON IS FROM THE H⁺ THAT HAS BEEN DONATED BY HYDROCHLORIC ACID

THESE ELECTRONS COME FROM THE OXYGEN IN WATER AND ARE USED TO FORM A BOND WITH THE INCOMING H+

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?

Your notes

- 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 aluminium oxide which reacts with both hydrochloric acid and sodium hydroxide:

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

$$Al_2O_3 \text{ (s)} + 2NaOH \text{ (aq)} + 3H_2O \text{ (l)} \rightarrow 2NaAl(OH)_4 \text{ (aq)}$$
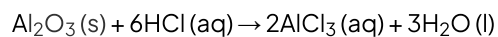
- When a compound is **amphiprotic**, it means it can act as a proton donor and as a proton acceptor
- Aluminium oxide is not amphiprotic, even though it is amphoteric

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

The pH Scale

Your notes

The pH Scale

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

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

  - 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 in 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

✏️ 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}
  - This means an increase of 2 pH units
  - Therefore, the final solution is **pH 3**

💡 Examiner 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

- 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 **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

The pH scale

Your notes
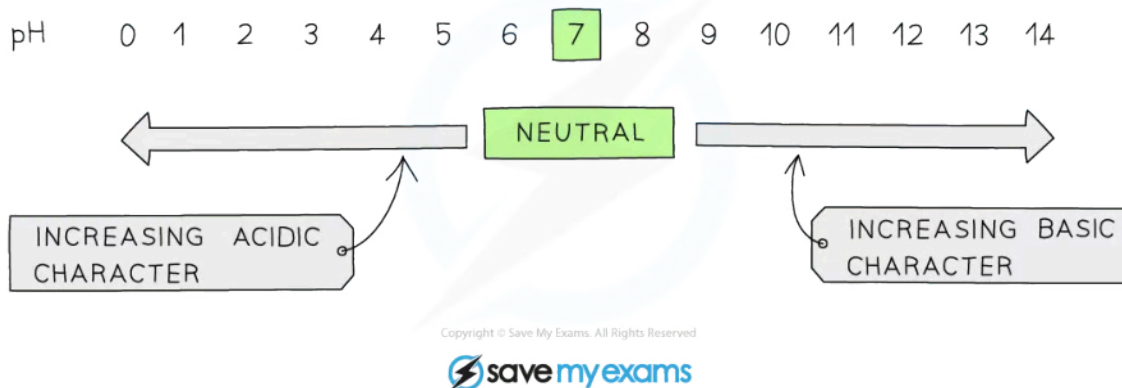
The pH scale showing acidity, neutrality and alkalinity

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

pH of water

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

$$pH = - \log_{10}[H^+ \text{(aq)}]$$

$$[H^+ \text{(aq)}] = \text{concentration of } H^+/H_3O^+ \text{ ions}$$

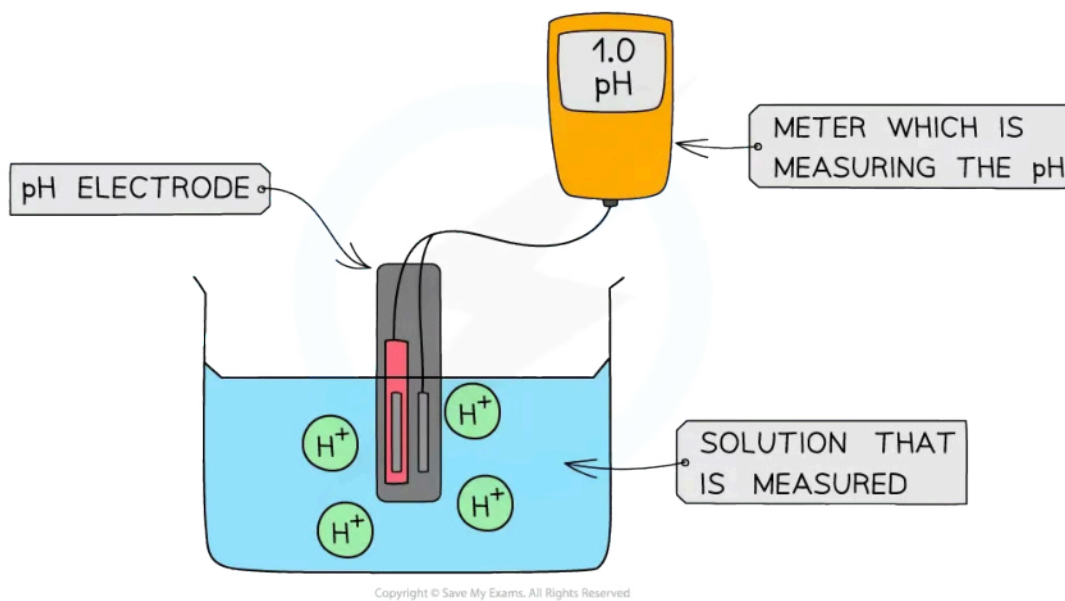$$pH = -\log(10^{-7}) = 7$$

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

How to measure pH

- 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

Using a pH meter

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

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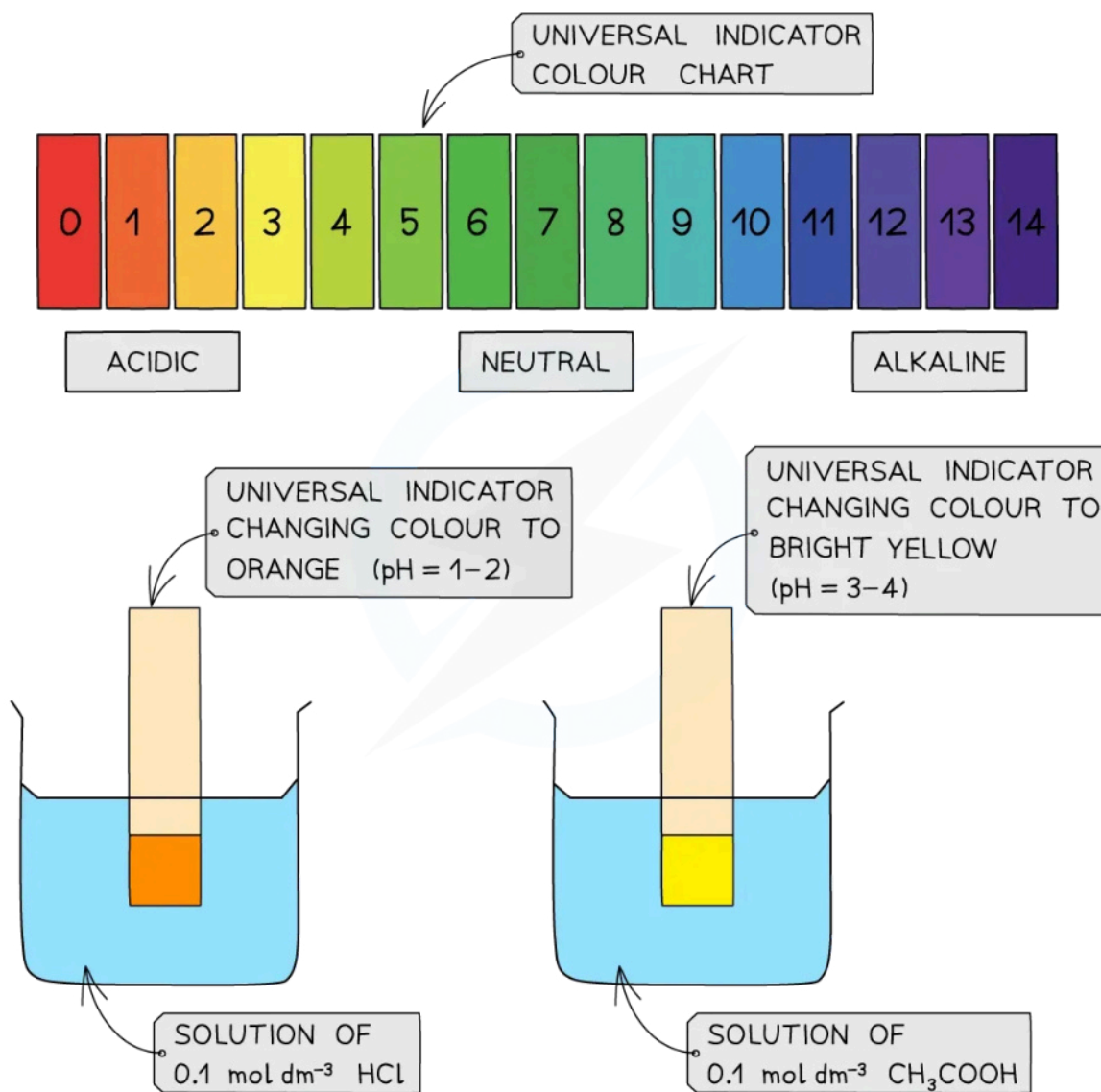
The diagram shows the change in colour of the universal indicator paper when dipped in a strong (HCl) and weak (CH_3COOH) acid. The colour chart is used to read off the corresponding pH values which are between 1–2 for HCl and 3–4 for CH_3COOH

The Ion Product of Water

The Ion Product of Water

pH of water

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

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

- The equilibrium constant for this reaction is:

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

$$K_c \text{ x } [H_2O] = [H^+][OH^-]$$

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

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

  - Where K_w (ion product of water) = K_c x $[H_2O]$ = 1.00×10^{-14} at 298K

- The product of the two ion concentrations is always 1.00×10^{-14}
- 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

> 🖊️ **Worked example**
>
> What is the pH of a solution of potassium hydroxide, KOH (aq) of concentration 1.0×10^{-3} mol dm^{-3} ?
>
> $K_w = 1.0 \times 10^{-14}$ at 298 K
>
> > **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**

How does temperature affect the ion product of water, K_w?

- The ionisation of water is an **endothermic** process

$$2H_2O \, (l) \rightleftharpoons H_3O^+ \, (aq) + OH^- \, (aq)$$

- In accordance with Le Châtelier's principle, an increase in temperature will result in the forward reaction being favoured
  - This causes an increase in the concentration of the hydrogen and hydroxide ions
  - This leads to the **magnitude of K_w increasing**
  - Therefore, the **pH will decrease**
- Increasing the temperature decreases the pH of water (becomes more acidic)
- Decreasing the temperature increases the pH of water (becomes more basic)

Graph to show how K_w changes with temperature
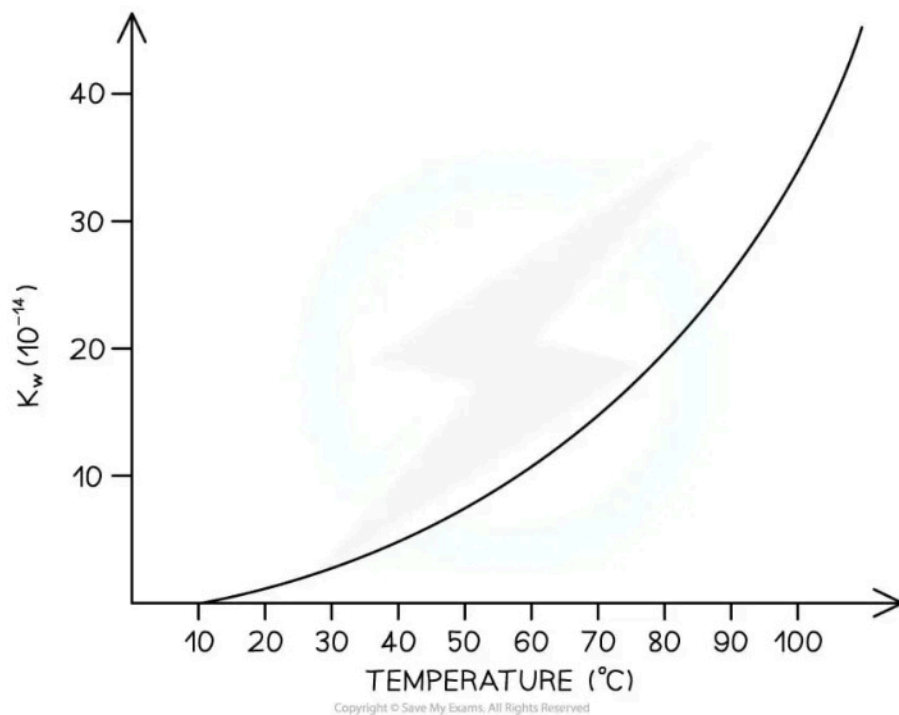
*As temperature increases, K_w increases so pH decreases*

Strong & Weak Acids

Strong & Weak Acids

Strong acids

- A **strong acid** is an acid that **dissociates** almost **completely** in aqueous solutions
  - Examples include 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

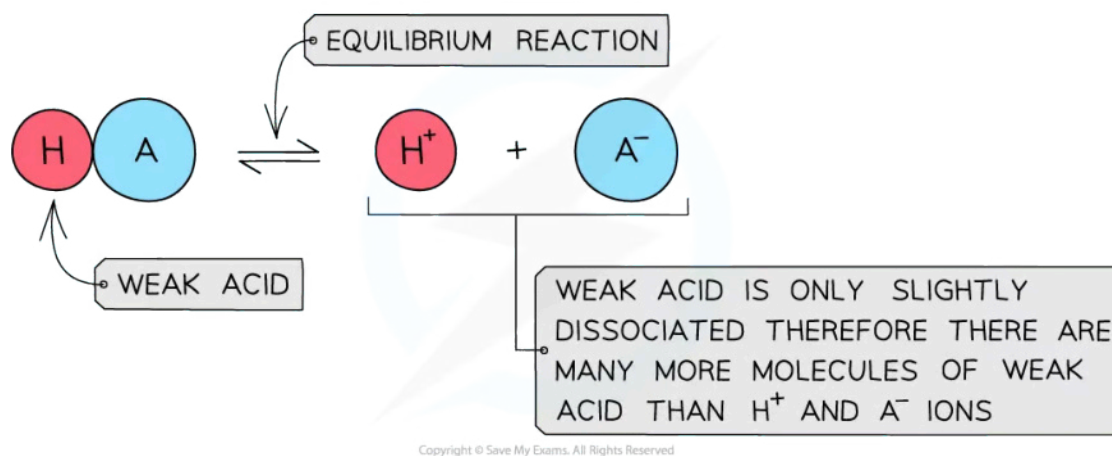#### Diagram to show the dissociation of a strong acid
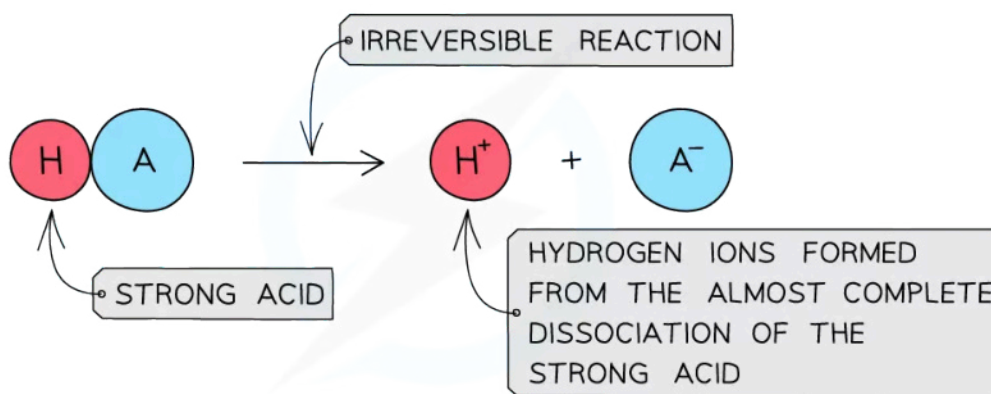
The diagram shows the complete dissociation of a strong acid in aqueous solution

- 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
  - $pH = -\log_{10}[H^+ (aq)]$
  - $[H+ (aq)]$ = concentration of H^+ / H_3O^+ ions
  - 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
  - E.g. 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 towards the **left** and an equilibrium is established

Diagram to show the dissociation of a weak acid

Your notes

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

Acid & Equilibrium Position Table

|	Strong Acids	Weak Acid
Position of equilibrium	Right	Left
Dissociation	Completely ($\rightarrow$)	Partially ($\rightleftharpoons$)
H^+ concentration	High	Low
pH	Use [strong acid] to calculate pH	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

- The strength of a Brønsted-Lowry acid depends on the ease with which it dissociates to release H^+ ions
  - This depends upon the strength of the bond that has to be broken to release H+

✏️
Your notes

- For example, for hydrogen halides, the size of the halogen atom increases in size going down Group 17 which increases the length of the H–X bond
- As longer bonds are weaker they need less energy to break
- The acid strength of the hydrogen halides increases down Group 17
  - HF < HCl < HBr < HI

Strong bases

- 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

Diagram to show the dissociation of a strong base

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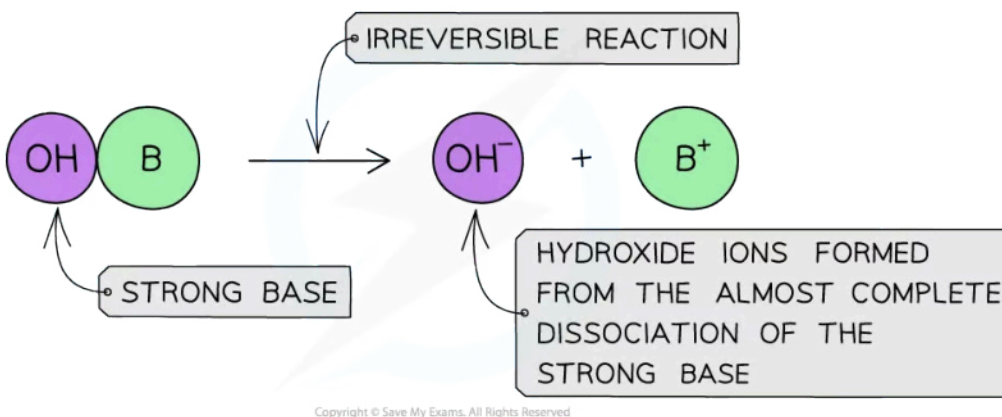
The diagram shows the complete dissociation of a strong base in aqueous solution

- 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
Diagram to show the dissociation of a weak base

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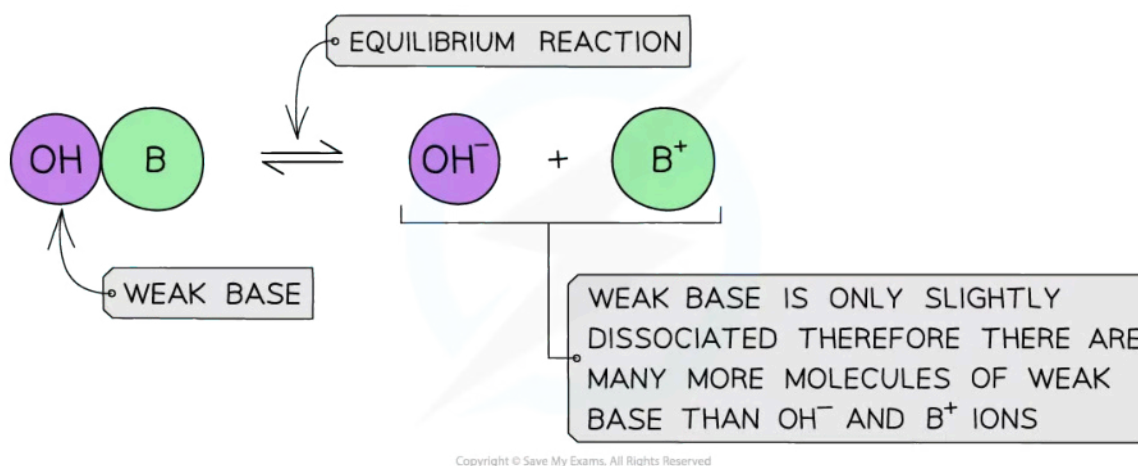
The diagram shows the partial dissociation of a weak base in aqueous solution

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

Base & Equilibrium Position Table

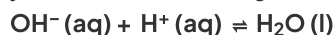|	Strong Base	Weak Base
Position of equilibrium	Right	Left
Dissociation	Completely ($\rightarrow$)	Partially ($\rightleftharpoons$)
OH^- concentration	High	Low
Examples	Group 1 metal hydroxides	NH_3 Amines Some transition metal hydroxides

Strength of conjugate acids and bases

- The conjugate base of HCl is the chloride ion, Cl^-,

- However, since the reverse reaction is virtually non-existent the chloride ion must be a very weak conjugate base

$$HCl\,(g) \rightarrow H^+\,(aq)\ +\ Cl^-\,(aq)$$
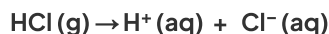
$$\text{acid} \qquad\qquad \text{conjugate base}$$

- In general, **strong acids** produce **weak conjugate bases** and **weak acids** produce **strong conjugate bases**
- A strong base is also fully ionised 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**

> 🔘 **Examiner 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: $\mathbf{H_2SO_4 \rightarrow H^+ + SO_4^-}$
>   - HSO_4^- acts as a weak acid: $\mathbf{HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}}$
>   - 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**

How to distinguish between strong and weak acid

- 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

$$HA \rightarrow H^+ + A^-$$
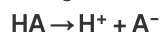
pH value of a Strong Acid & Weak Acid Table

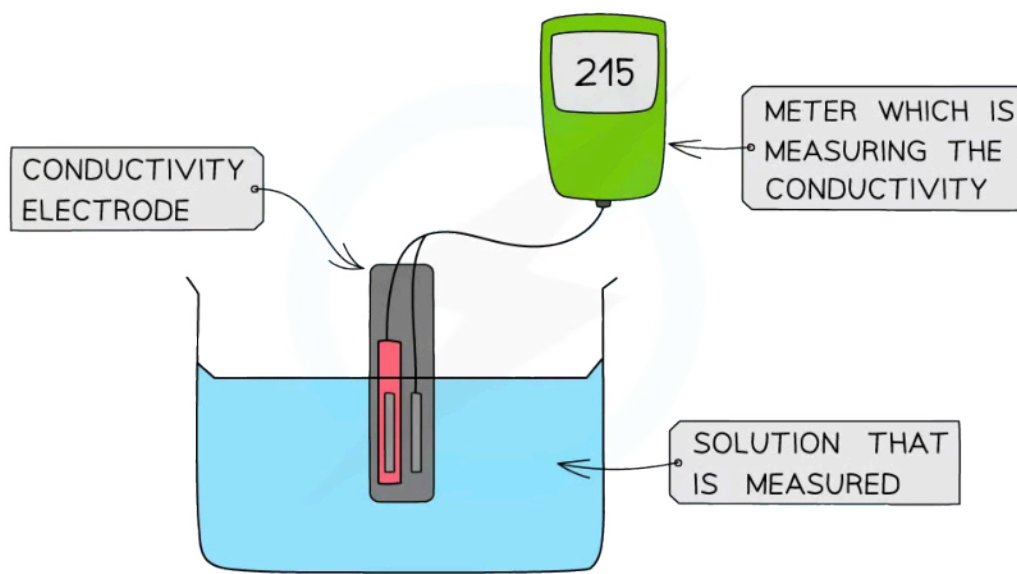| Acid	pH of 0.1 mol dm^{-3} solution

HCl (strong)	1
CH_3COOH (weak)	2.9

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

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

**Diagram to show how to measure the electrical conductivity of an acid**

A digital conductivity meter 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

Diagram to show how a strong acid reacts with magnesium
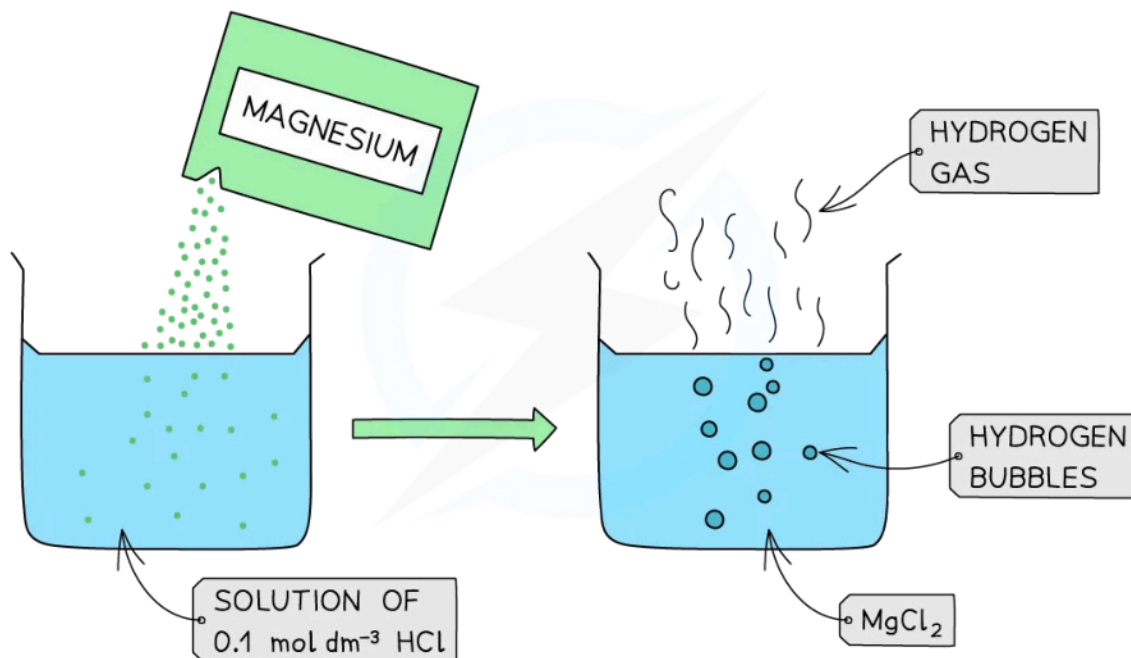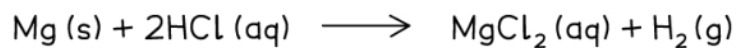
Your notes

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

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

Diagram to show how a weak acid reacts with magnesium

Your notes

$$Mg\,(s) + 2CH_3COOH\,(aq) \longrightarrow Mg(CH_3COO)_2\,(aq) + H_2\,(g)$$
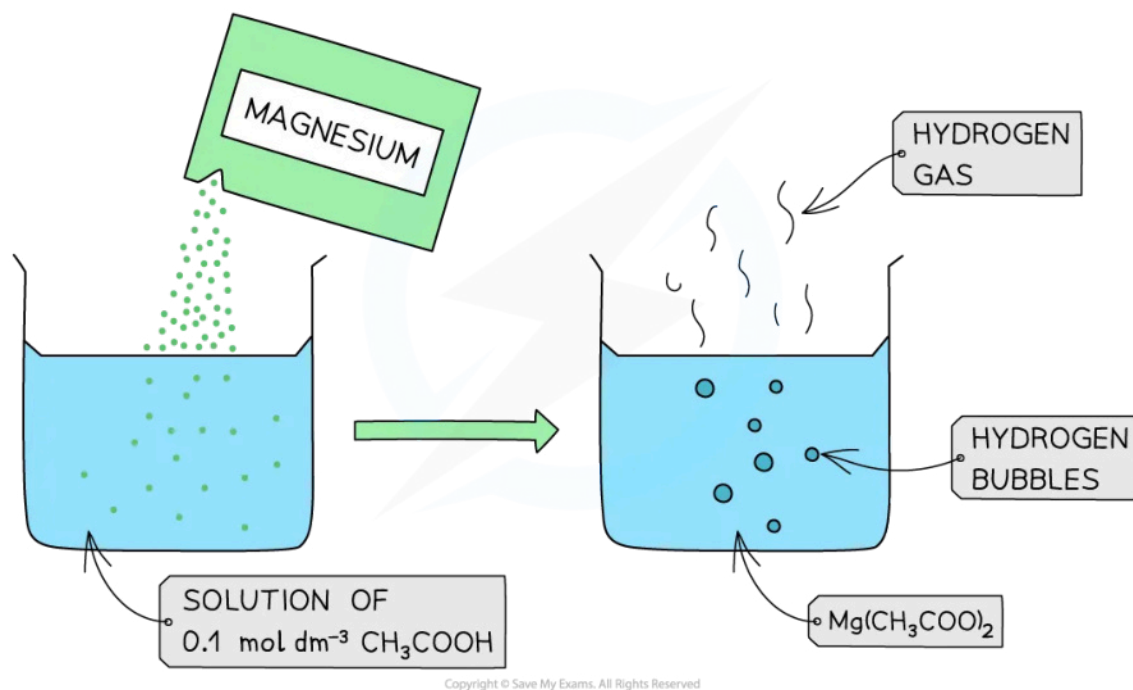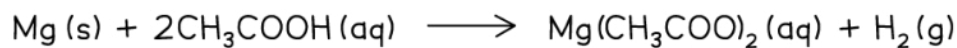
The diagram shows the reaction of $0.1\,mol\,dm^{-3}$ of a weak acid (CH_3COOH) with Mg. The reaction produces fewer bubbles of hydrogen gas due to the lower concentration of H^+ present in solution

- 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**

💡 **Examiner Tip**

- The above-mentioned properties of strong and weak acids depend on their ability to dissociate and form H^+ ions
- Stronger acids dissociate more
  - This means that they produce a greater concentration of H^+ ions resulting in:
    - Lower pH values
    - Greater electrical conductivity
    - More vigorous reactions with reactive metals.

Neutralisation Reactions

Neutralisation Reactions

- A neutralization 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:

acid + base (alkali) → salt + water

- The proton of the acid reacts with the hydroxide of the base to form water:
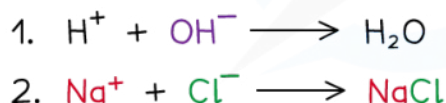
$H^+ (aq) + OH^- → H_2O (l)$

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

Diagram to show neutralisation between an acid and a base

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

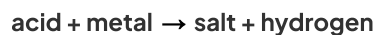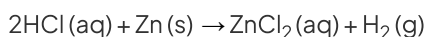| Acid reacted	Salt produced
Hydrochloric acid	A chloride
Nitric acid	A nitrate
Sulfuric acid	A sulfate

> 💡 **Examiner 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}

Metals and acids

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

$$\textbf{acid + metal} \rightarrow \textbf{salt + hydrogen}$$

- For example:

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

$$\text{hydrochloric acid + zinc} \rightarrow \text{zinc chloride + hydrogen}$$
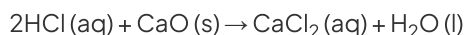
- Clearly, the extent of the reaction depends on the **reactivity** of the metal and the **strength** of the acid
- Very reactive metals would react dangerously with acids and these reactions are not usually carried out
- Metals low in **reactivity** do not react at all
  - For instance, copper does not react with dilute acids
- **Stronger acids** will react **more vigorously** with metals than weak acid
- 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

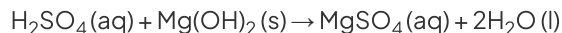- The reaction of an acid with a metal oxide forms two products:

$$\textbf{acid + metal oxide} \rightarrow \textbf{salt + water}$$

- For example:

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

$$\text{hydrochloric acid + calcium oxide} \rightarrow \text{calcium chloride + water}$$

Metals and hydroxides

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

$$\textbf{acid + metal hydroxide} \rightarrow \textbf{salt + water}$$

- A suitable example might be:

$$H_2SO_4 \, (aq) + Mg(OH)_2 \, (s) \rightarrow MgSO_4 \, (aq) + 2H_2O \, (l)$$
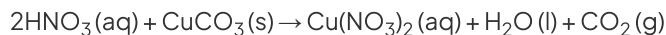
sulfuric acid + magnesium hydroxide → magnesium sulfate + water

Metals and carbonates

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

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

> 💡 **Examiner 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

- 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 hydrogen carbonate
Nitrates	NO_3^-	nitric	HNO_3	
Chlorides	Cl^-	hydrochloric	HCl	
Ethanoates	CH_3COO^-	ethanoic	CH_3COOH	
Ammonium	NH_4^+	any	–	aqueous ammonia

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

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🖊 **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 react with acids to produce a salt and water as the only products

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Your notes

pH Curves

pH Curves

Strong acid – strong base pH curve

- During a titration, a pH meter can be used and a pH curve plotted
- A pH curve is a graph showing how the pH of a solution changes as the acid (or base) is added in a strong acid - strong base titration, e.g.
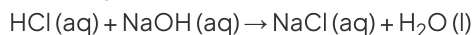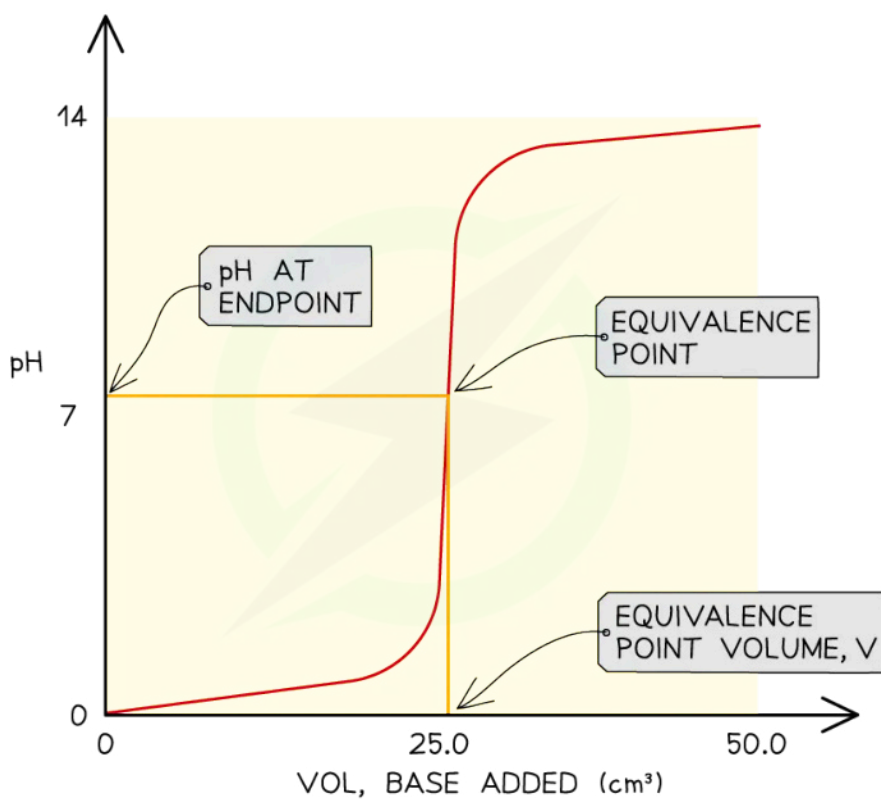
$$HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l)$$

Diagram to show the general characteristics of a strong acid–strong base pH curve



The characteristics of a pH curve

- All pH curves show an s-shape curve
- pH curves yield useful information about how the acid and alkali react together with stoichiometric information
- The midpoint of the inflection is called the **equivalence** or **stoichiometric point**

Your notes
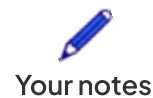
- From the curves you can:
  - Determine the pH of the acid by looking where the curve starts on the y-axis
  - Find the pH at the equivalence point
  - Find volume of base at the equivalence point
  - Obtain the range of pH at the vertical section of the curve

How to calculate the pH depending on the volume of base added

- If base is added to the conical flask then the pH of the solution will rise during the titration
- Let's look at the reaction between 50 cm^3 of 0.10 mol dm^{-3} HCl (aq) and 50 cm^3 of 0.10 mol dm^{-3} of NaOH (aq)

$$HCl \text{ (aq)} + NaOH \text{ (aq)} \rightarrow NaCl \text{ (aq)} + H_2O \text{ (l)}$$
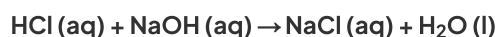
- At the **start:**
  - At the start of the titration, the conical flask will only contain a strong acid so the pH can be calculated by
    - pH = $-log_{10}$[H$^+$]
    - pH = $-log_{10}$[0.10] = 1.0

- After **25.00 cm^3** of **NaOH** has been added
  - Now, we must consider what is in excess
  - There is more acid in the flask than base in terms of volume, some of the acid has been neutralised, so we must calculate the excess moles of one of the reactants using $n = c$ (mol dm^{-3}) x v (dm^3)
    - n(HCl) = 0.10 x 0.050 = 0.0050 mol
    - n(NaOH) = 0.10 x 0.025 = 0.0025 mol
    - n(Excess HCl) = 0.0050 – 0.0025 = 0.00250 mol
      - New volume = 0.0750 dm^3
    - [H$^+$] = $\dfrac{0.0025}{0.0750}$ = 0.0333 mol dm^{-3}
    - so pH = 1.5

- After **49.00 cm^3** of **NaOH** has been added
  - n(HCl) = 0.10 x 0.050 = 0.0050 mol
  - n(NaOH) = 0.10 x 0.049 = 0.0049 mol
  - n(Excess HCl) = 0.0050 – 0.0049 = 0.0001 mol
    - New volume = 0.0990 dm^3
  - [H$^+$] = $\dfrac{0.0001}{0.0990}$ = 0.00101 mol dm^{-3}
  - so pH = 3.0

- After **50.00 cm^3** of **NaOH** has been added the acid has been completely neutralised by the base, so the solution only contains NaCl and H$_2$O, therefore the pH = 7.0
- After **51.00 cm^3** of **NaOH** has been added
  - n(Added NaOH) = 0.10 x 0.051 = 0.0051 mol
  - n(Excess NaOH) = 0.0051 – 0.0050 = 0.0001 mol

- New volume = 0.101 dm^3
- $[OH^-] = \dfrac{0.0001}{0.101} = 0.00099$ mol dm^{-3}
- p**OH** = 3.0
- so pH = 11.0