Cambridge (CIE) A Level Chemistry

Your notes

Brønsted–Lowry Theory of Acids & Bases

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* Acids & Bases
* Brønsted–Lowry Theory
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Your notes

Common Acids

- An **acid** is a substance that **neutralises** a base forming a **salt** and **water**:

$$\text{acid + base} \rightleftharpoons \text{salt + water}$$

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

- Acids are also substances that release **hydrogen ions** when they dissolve in water:

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

Acid dissociation

Acids dissociate in water to release a hydrogen ion

Common acids

- **Hydrochloric acid**
  - HCl forms $H^+ + Cl^-$ in water
- **Nitric acid**
  - HNO_3 forms $H^+ + NO_3^-$ in water
- **Sulfuric acid**
  - H_2SO_4 forms $H^+ + SO_4^{2-}$ in water
- **Ethanoic acid**
  - CH_3COOH forms $H^+ + CH_3COO^-$ in water
- Monoprotic inorganic acids, such as hydrochloric acid, fully dissociate into their ions
- Organic acids, such as **carboxylic acids**, do not fully dissociate into their ions
  - Only some of the hydrogen atoms can form ions

Common Alkalis

- A **base** is a compound that **neutralises** an acid forming a **salt** and **water**

$$\text{acid + base} \rightleftharpoons \text{salt + water}$$

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

- A base is a substance that **accepts** hydrogen ions or a compound that contains **oxide** or **hydroxide** ions
  - For example, when the base ammonia is added to water, the ammonium ion and hydroxide ions are formed:

$$NH_3\,(g) + H_2O\,(l) \rightarrow NH_4^+\,(aq) + OH^-\,(aq)$$

- For example, when sodium hydroxide is dissolved in solution, sodium ions and hydroxide ions are formed:

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

- A base that is **soluble** in water is called an **alkali**

Common alkalis table

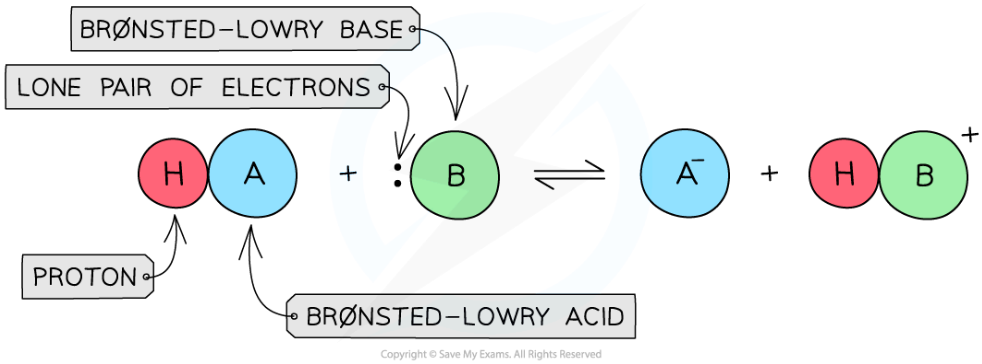
- **Sodium hydroxide**
  - $NaOH$ forms $Na^+ + OH^-$ in water
- **Potassium hydroxide**
  - KOH forms $K^+ + OH^-$ in water
- **Aqueous ammonia**
  - NH_3 forms $NH_4^+ + OH^-$ in water

Your notes

Brønsted–Lowry Theory

- **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**

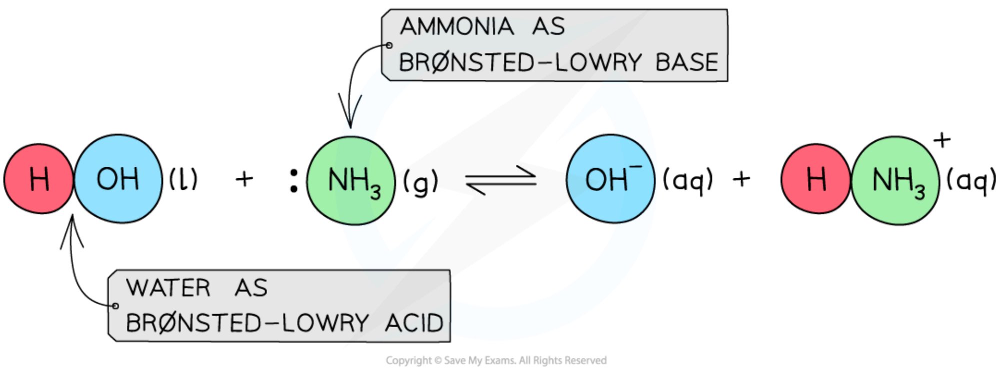
How an acid acts as a Brønsted–Lowry proton donor



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

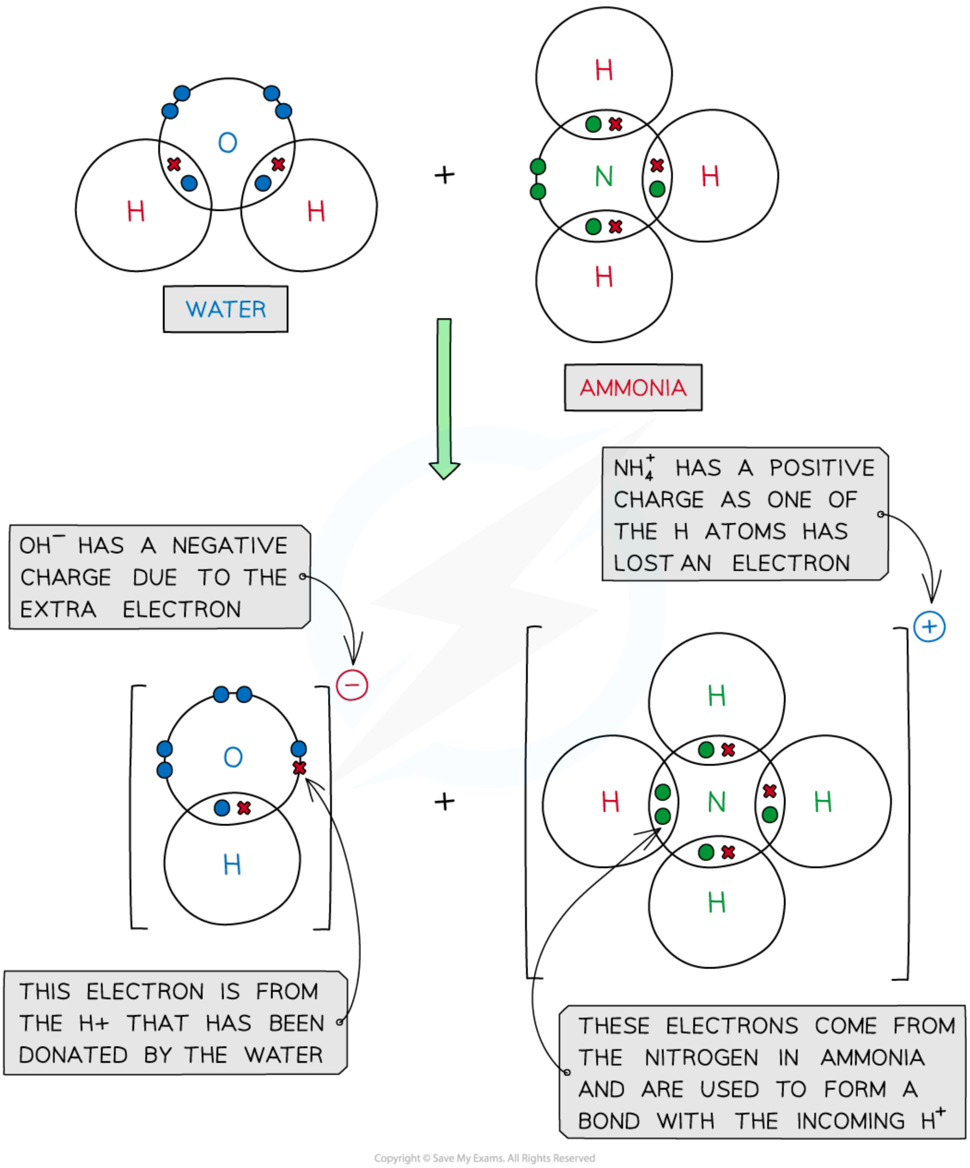
- Species that can act both as acids and bases are called **amphoteric**
  - Eg. water as a Brønsted-Lowry acid

Water acting as a Brønsted–Lowry acid



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

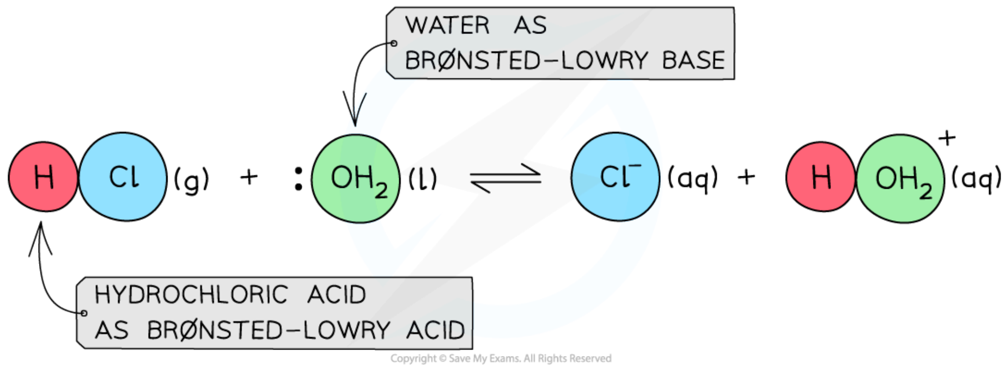
Dot and cross diagram showing the Brønsted-Lowry behaviour of water with ammonia



The diagram shows a dot & cross 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

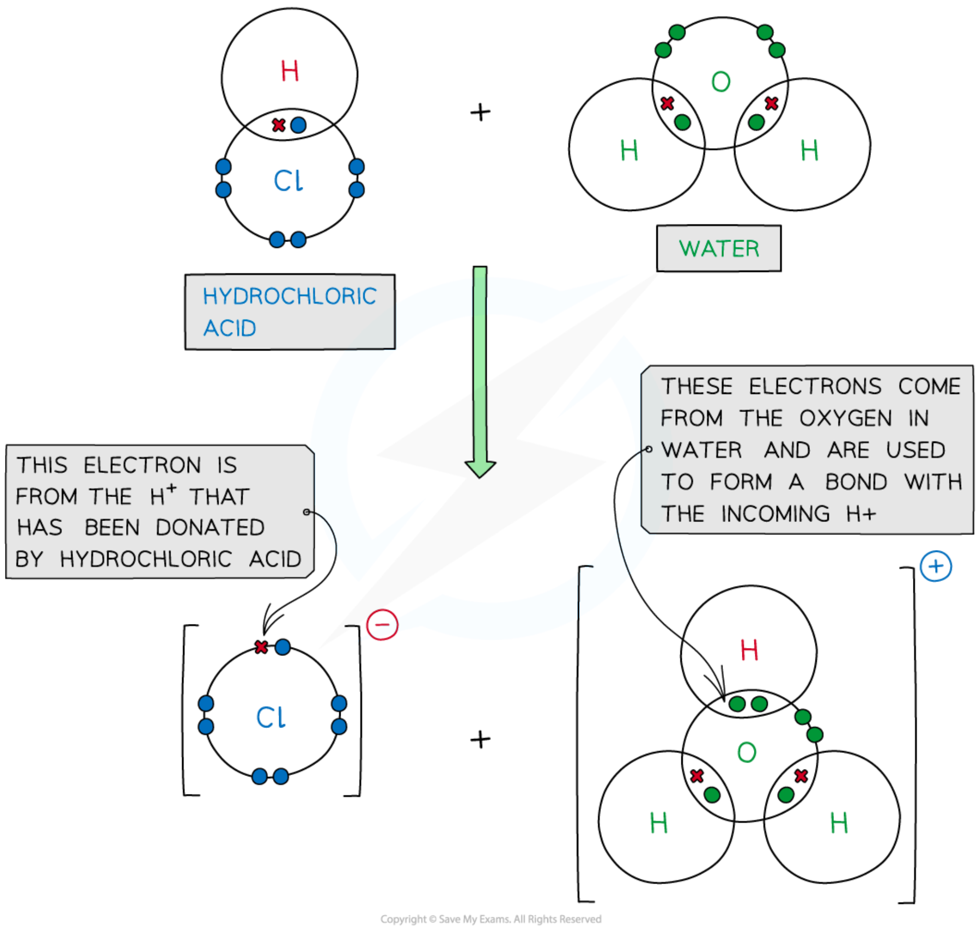
- E.g. water as a Brønsted-Lowry base

Water acting as a Brønsted-Lowry base



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

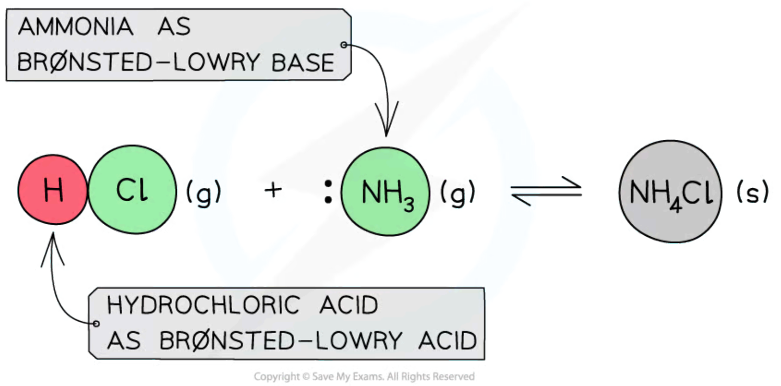
Dot and cross diagram showing the Brønsted-Lowry behaviour of water with hydrochloric acid



The diagram shows a dot & cross 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

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

A Brønsted-Lowry acid and base reaction



HCl acts as a Brønsted-Lowry acid by donating a proton while ammonia acts as a Brønsted-Lowry base by accepting a proton

Examiner Tips and Tricks

- 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

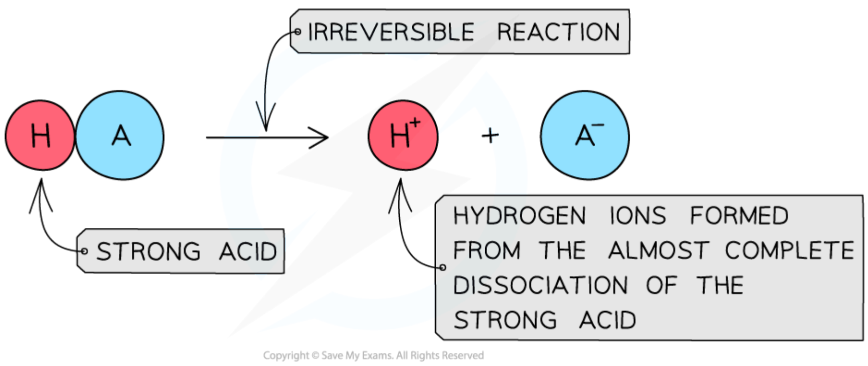
Your notes

Acid & Base Dissociation

Strong acids

- A **strong acid** is an acid that **dissociates** almost **completely** in aqueous solutions
  - E.g. 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 showing the dissociation of a strong acid in aqueous solution



In an aqueous solution, a strong acid almost completely dissociates

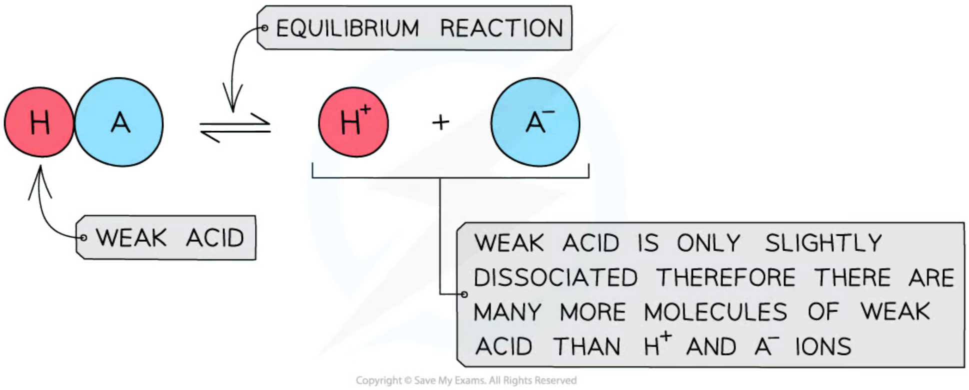
- The solution formed is **highly acidic** due to the high concentration of the H^+ / H_3O^+ ions
- Since the **pH** depends on the concentration of H^+ / H_3O^+ ions, the pH can be calculated if the concentration of the strong acid is known
  - The concentration of H^+ / H_3O^+ ions can be written as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

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

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 to the **left** and an equilibrium is established

Diagram showing the dissociation of a weak acid in aqueous solution



In an aqueous solution, a weak acid does not fully dissociate

- The solution is **less acidic** due to the lower concentration of H^+ / H_3O^+ ions
- Finding the **pH** of a weak acid is a bit more complicated as now the concentration of H^+ ions is not equal to the concentration of acid
- To find the concentration of H^+ ions, the acid dissociation constant (K_a) should be used

Acid & equilibrium position summary

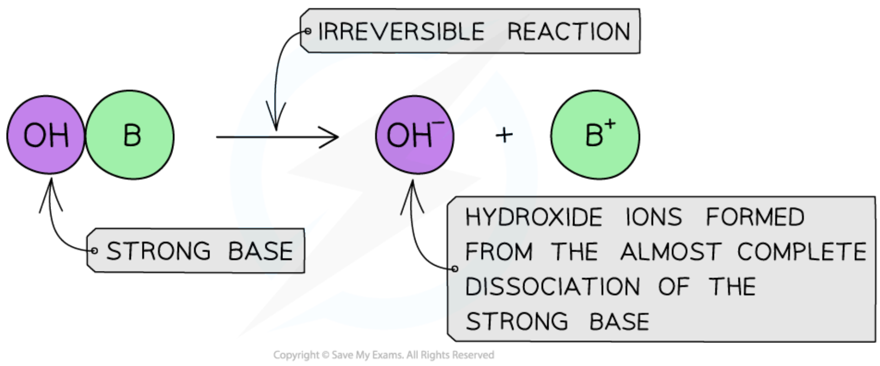
- Position of equilibrium
  - Strong acid; right
  - Weak acid; left
- Dissociation
  - Strong acid; fully dissociated (→)
  - Weak acid; partially dissociated ($\rightleftharpoons$)
- H^+ concentration
  - Strong acid; high concentration
  - Weak acid; low concentration
- pH
  - Strong acid; use [strong acid] for $[H^+]$
  - Weak acid; use K_a to find $[H^+]$
- Examples
  - Strong acid; HCl, HNO_3, H_2SO_4 (first ionisation)
  - Weak acid; Organic acids, e.g. ethanoic acid, HCN, H_2S, H_2CO_3

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 showing the dissociation of a strong base in aqueous solution



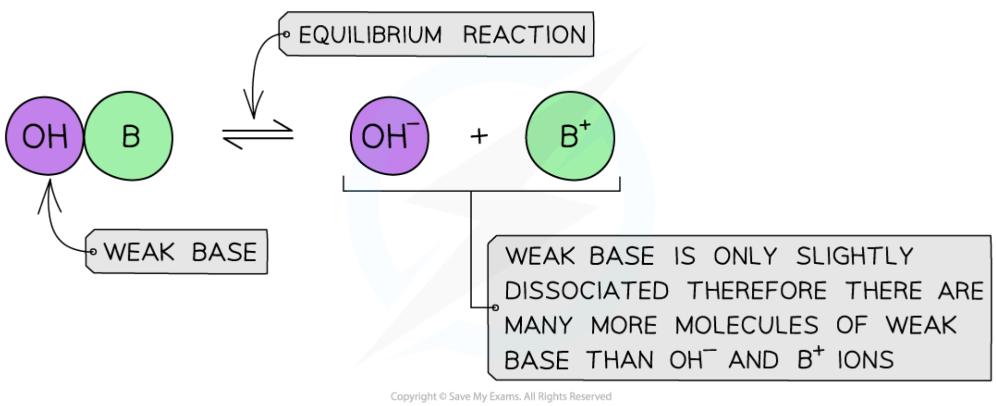
In an aqueous solution, a strong base almost completely dissociates

- 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
  - E.g. 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 showing the dissociation of a weak base in aqueous solution



In an aqueous solution, a weak base does not fully dissociate

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

Base & equilibrium position summary

- Position of equilibrium

Your notes

- Strong base; right
- Weak base; left
- Dissociation
    - Strong base; fully dissociated ($\rightarrow$)
    - Weak base; partially dissociated ($\rightleftharpoons$)
- OH^- concentration
    - Strong base; high concentration
    - Weak base; low concentration
- Examples
    - Strong base; Group 1 metal hydroxides
    - Weak base; NH_3 amines, some transition metal hydroxides

Your notes

Examiner Tips and Tricks

- 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:

$$\mathbf{HCl\ (g) \rightarrow H^+\ (aq) + Cl^-\ (aq)}$$

or

$$\mathbf{HCl\ (g) + H_2O\ (l) \rightarrow H_3O^+\ (aq) + Cl^-\ (aq)}$$

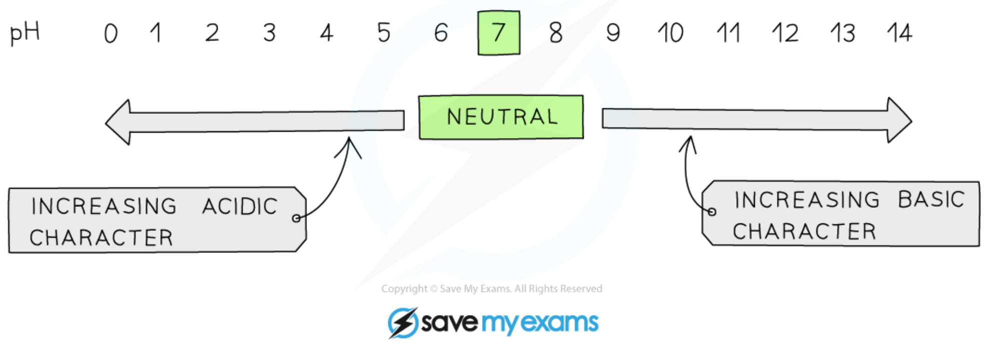
- Remember that some acids are both strong and weak acids – for example, H_2SO_4 (sulfuric acid) has two hydrogen ions that can ionise.
    - H_2SO_4 acts as a strong acid: $\mathbf{H_2SO_4 \rightarrow H^+ + HSO_4^-}$
    - HSO_4^- acts as a weak acid: $\mathbf{HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}}$
- 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.**

pH Scale

The pH Scale

- The pH scale is a numerical scale that shows how **acidic** or **alkaline** a solution is
- The values on the pH scale go from 1–14 (extremely acidic substances have values of below 1)
  - All acids have pH values **below** 7
  - All alkalis have pH values **above** 7
- The **lower** the pH, the **more acidic** the solution is
- The **higher** the pH, the **more alkaline** the solution is

The pH scale



The pH scale shows the acidity, neutrality and alkalinity of chemicals

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 expression for this reaction is:

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

- The equilibrium constant expression can be rearranged to:

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

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

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

- Where K_w **(ionic product of water) = K_c x $[H_2O]$ = 10^{-14} mol^2 dm^{-3} at 298K**
- 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:

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

- Where $[H^+\,(aq)]$ is the concentration of H^+ / H_3O^+ ions
- So, the calculation is:
  - **pH = $-\log\,(10^{-7}) = 7$**
- Thus, water has a pH of 7

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

Your notes

Your notes

Strong & Weak Acids & Bases

- 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

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

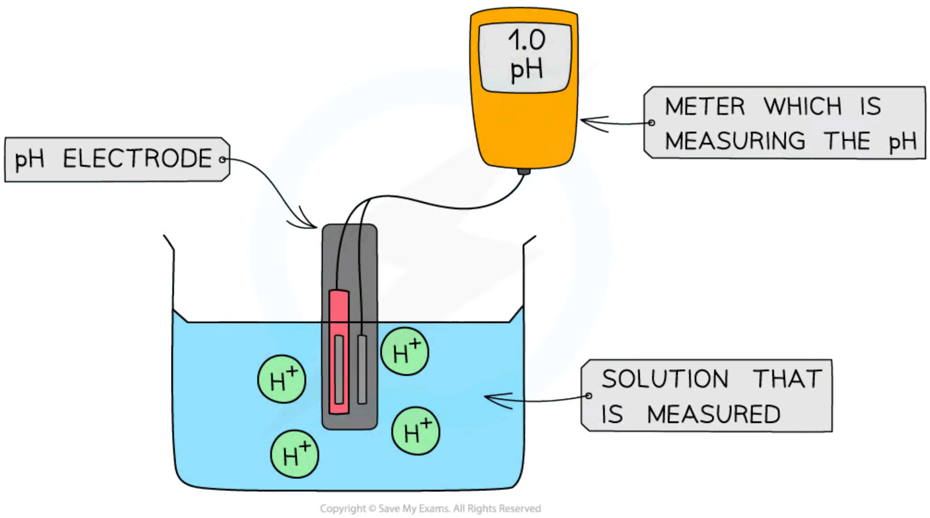
$$HA \rightarrow H^+ + A^-$$

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

pH values of a strong & weak acids

- pH of 0.1 mol dm^{-3} solution:
  - HCl (strong); pH 1
  - CH_3COOH (weak); pH 2.0
- 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 digital pH meter

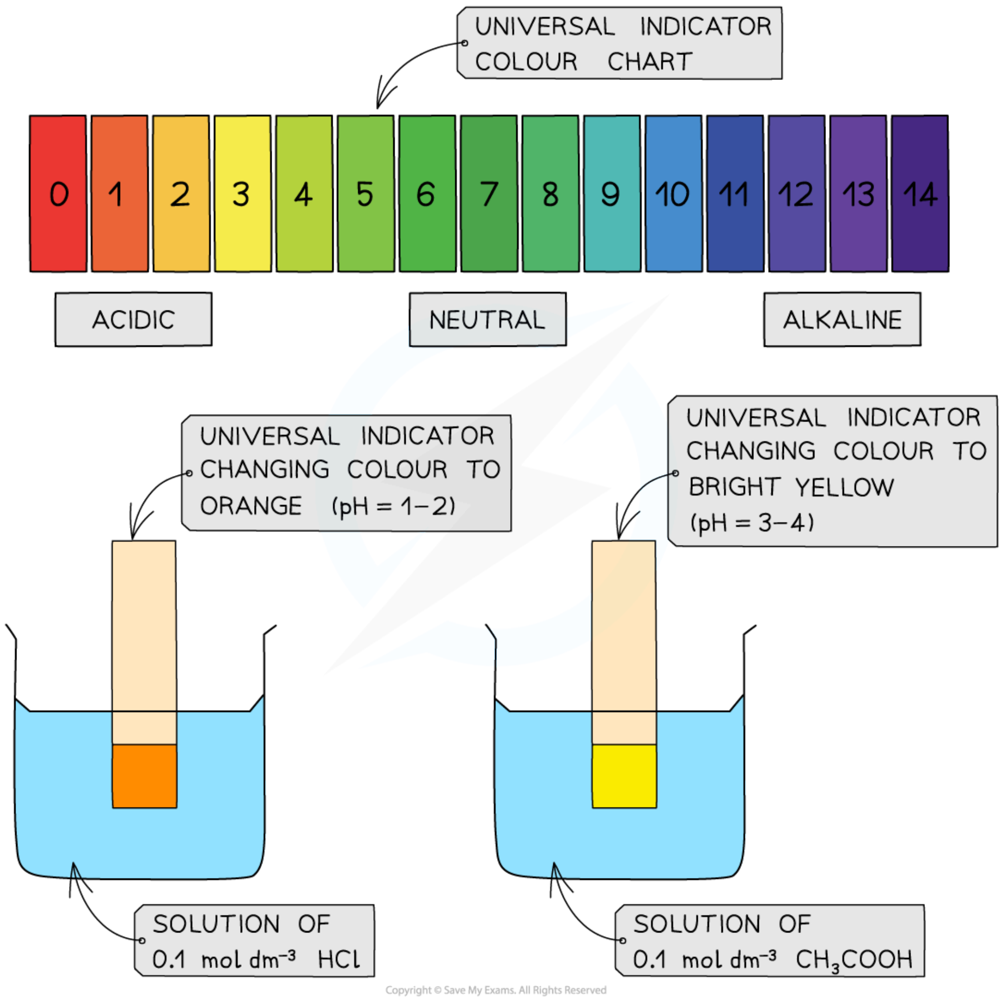


The diagram shows a digital pH meter measures the pH of a solution using a pH electrode

- 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

Your notes

How to use universal indicator paper



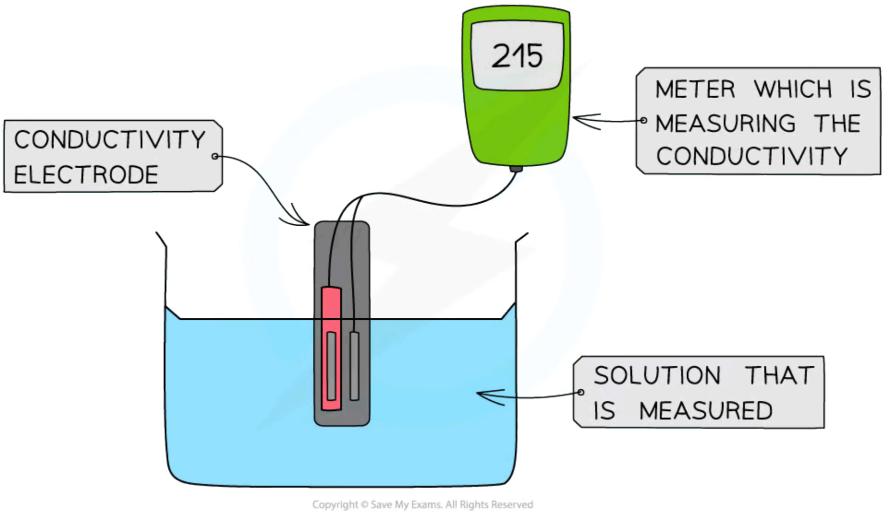
The diagram shows the change in colour of the universal indicator paper when dipped in a strong and weak acid. The colour chart is used to read off the corresponding pH values which are between 1–2 for a strong acid and 3–4 for a weak acid

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

Using a digital conductivity meter

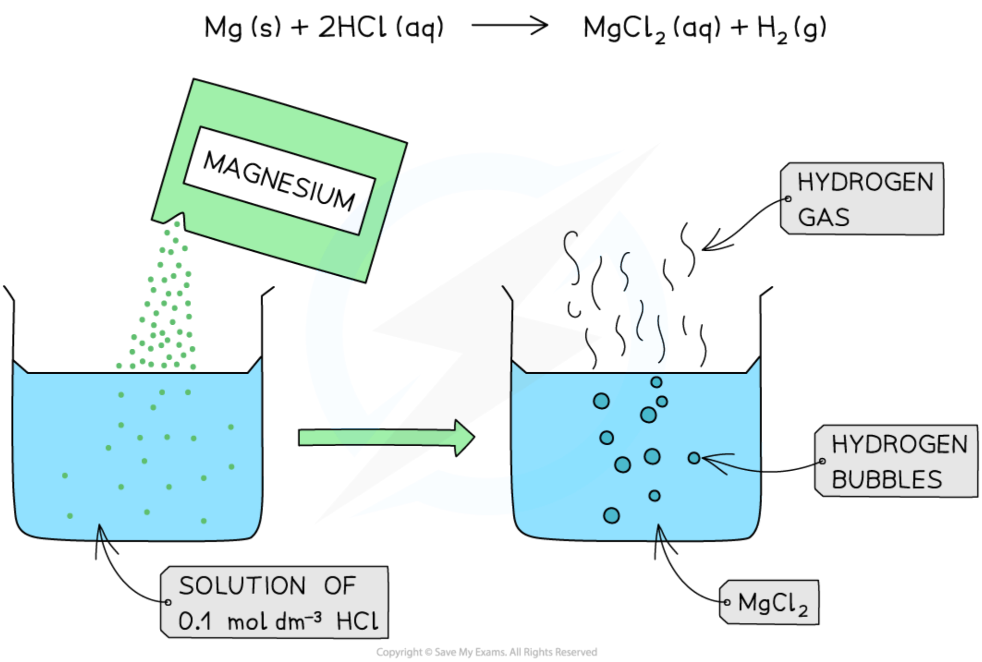


The diagram shows a digital conductivity meter that measures the electrical conductivity of a solution using an electrode

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

The reaction of 0.1 mol dm^{-3} of a strong acid, HCl, with Mg

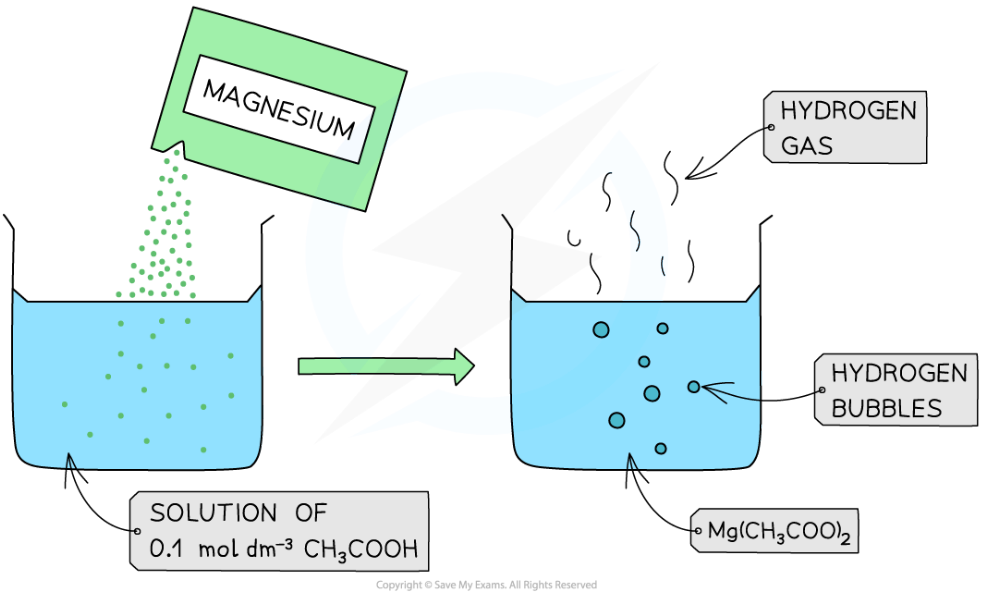


The reaction produces a lot of bubbles and hydrogen gas due to the high concentration of H+ present in the solution

Your notes

The reaction of 0.1 mol dm^{-3} of a weak acid, CH_3COOH, with Mg

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



The reaction produces fewer bubbles and hydrogen gas due to the lower concentration of H+ present in the solution

Examiner Tips and Tricks

- 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.

Neutralisation Reactions

- 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:

acid + base (alkali) → salt + water

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

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

- The spectator ions which are not involved in the formation of water are Na^+ (aq) + Cl^- (aq)
  - These react to form the salt:

$$\mathbf{Na^+ (aq) + Cl^- (aq) \rightarrow NaCl (aq)}$$

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

Salts produced from certain acids

- Hydrochloric acid forms **chloride** salts
- Sulfuric acid forms **sulfate** salts
- Nitric acid forms **nitrate** salts
- Ethanoic acid **ethanoate** salts

Examiner Tips and Tricks

Note that the reaction of an acid and metal carbonate also forms carbon dioxide:

acid + metal carbonate → salt + water + carbon dioxide

Your notes

pH Titration Curves

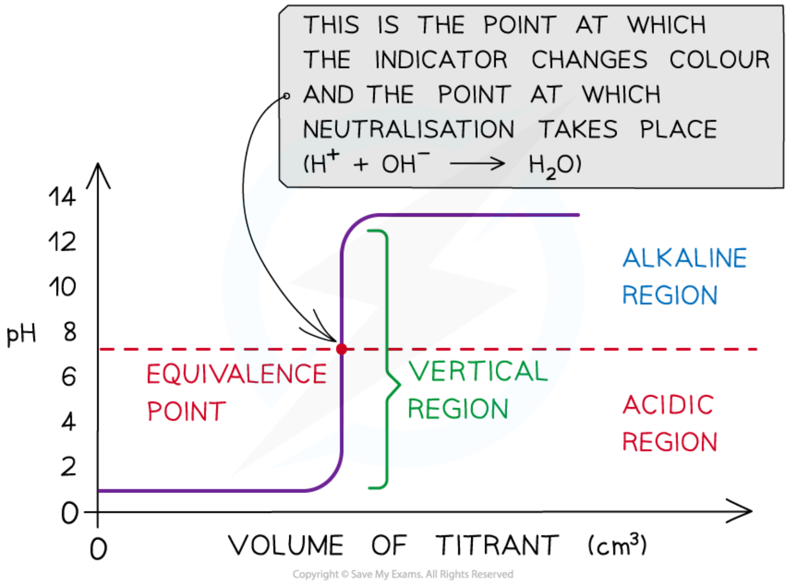
What are pH titration curves?

- Titration is a technique used in neutralisation reactions between acids and alkalis to determine the concentration of the unknown solution
- It involves adding **a titrant** of known concentration from a burette into a conical flask containing the **analyte** of unknown concentration
- An indicator is added which will change colour at the **endpoint** of the titration
- The endpoint is the point at which an equal number of moles of **titrant** and **analyte** react with each other
- The equivalence point is halfway along the **vertical region** of the curve

Equivalence point → moles of alkali = moles of acid

- This is also known as the **equivalence point** and this is the point at which **neutralisation** takes place

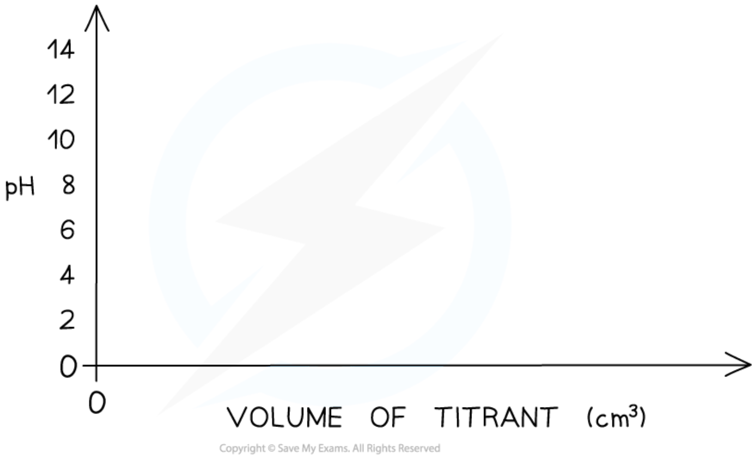
Example pH titration curve



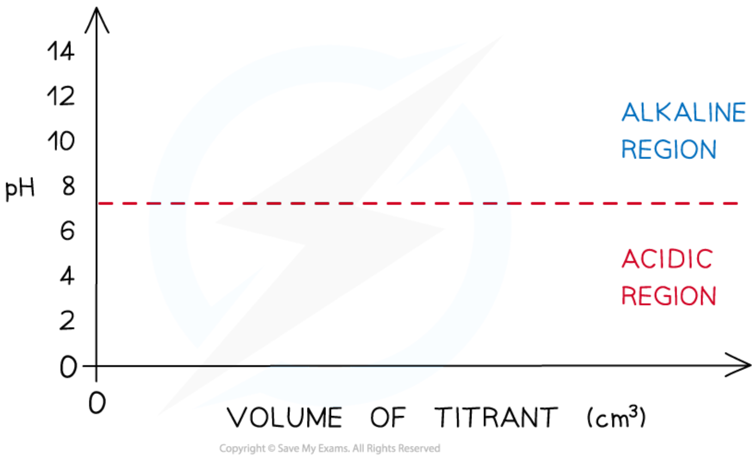
The equivalence point is the point at which an equal number of moles of titrant and analyte have reacted

Sketching a pH titration curve

- Draw axes with volume added (cm^3) on the x-axis and pH on the y-axis

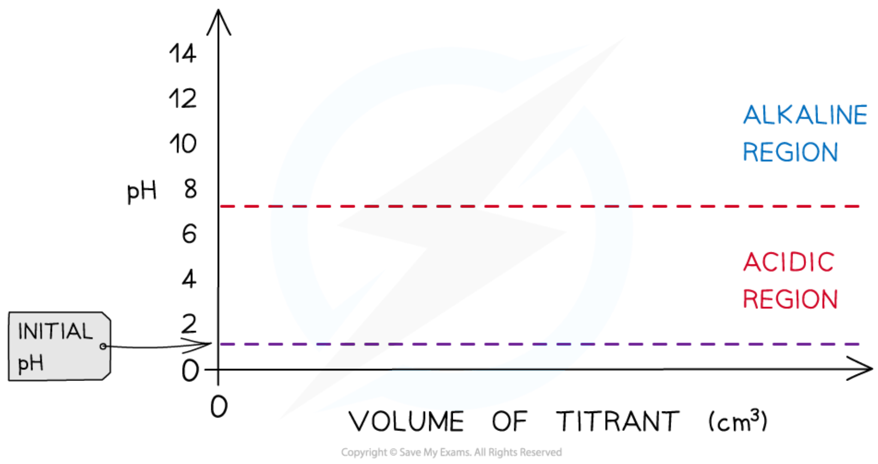


- Draw a horizontal line running parallel to the x-axis at pH 7
  - Everything below this line will be in the acidic **region** and everything above it in the **alkaline** region

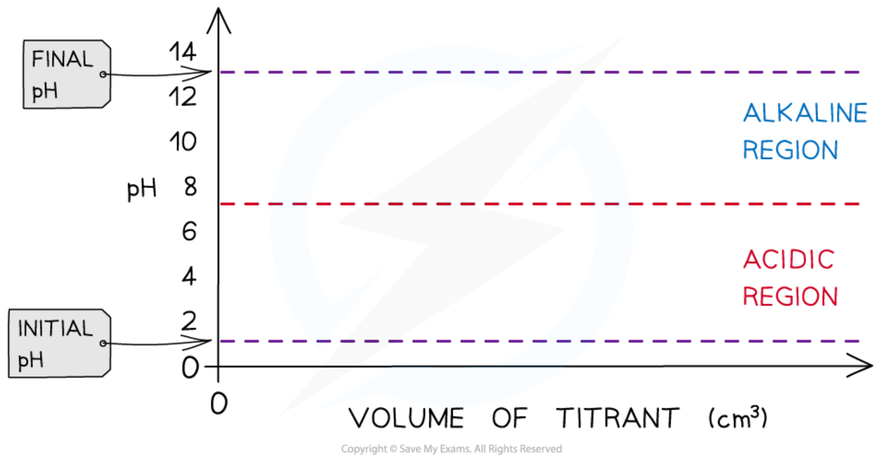


- Determine which substance is in the conical flask
  - If it is **a strong acid** the initial pH is about 1 or 2
  - If it is **a weak acid t**he initial pH is about 2–3
  - If it is **a strong alkali** the initial pH is about 13–14
  - If it is **a weak alkali** the initial pH is about 11

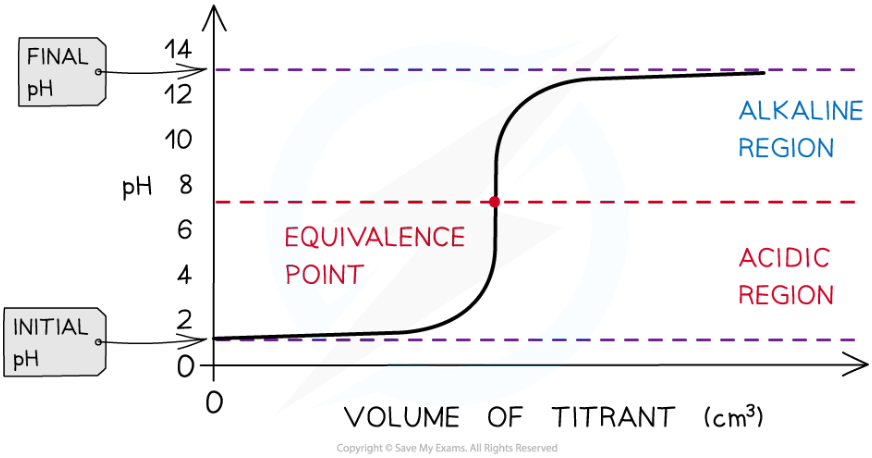
Your notes



- Determine what type of acid and alkali are used:
  - Strong acid + strong alkali
  - Strong acid + weak alkali
  - Weak acid + strong alkali
  - Weak acid + weak alkali



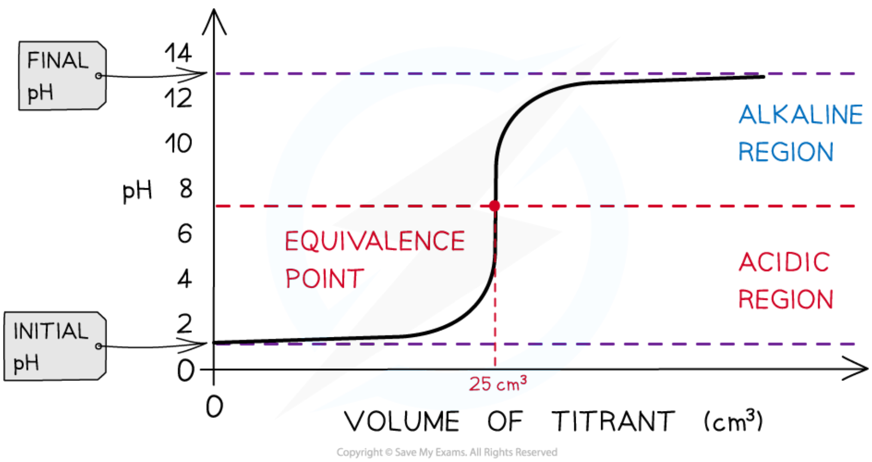
- Draw the pH titration curve



Strong acid + strong alkali pH titration curve

- Initially, there are only H^+ ions present in the solution from the dissociation of the strong acid (HCl) (initial pH about 1–2)
- As the **volume of strong alkali (NaOH) added increases**, the pH of the HCl solution slightly increases too as more and more H^+ ions react with the OH^- from the NaOH to form water
- The change in pH is not that much until the volume added gets close to the equivalence point
- The pH surges upwards very steeply
- The **equivalence point** is the point at which all H^+ ions have been neutralised
  - Therefore, the pH is 7 at the equivalence point
- Adding more NaOH will increase the pH as now there is an **excess in OH^- ions** (final pH about 13–14)

pH titration curve for a strong acid + strong alkali

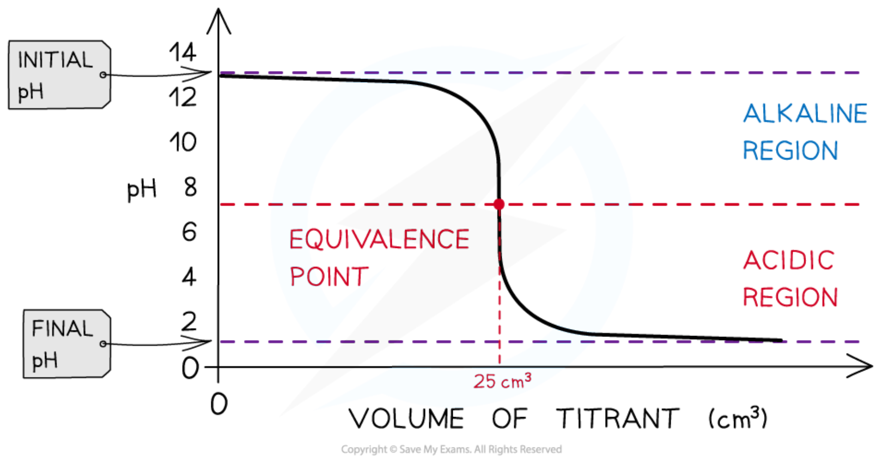


The diagram shows a pH titration curve of hydrochloric acid with sodium hydroxide

Your notes

- The pH titration curve for HCl added to a NaOH has the **same shape**
- The initial pH and final pH are the other way around
- The equivalence point is still 7

pH titration curve for a strong alkali + strong acid

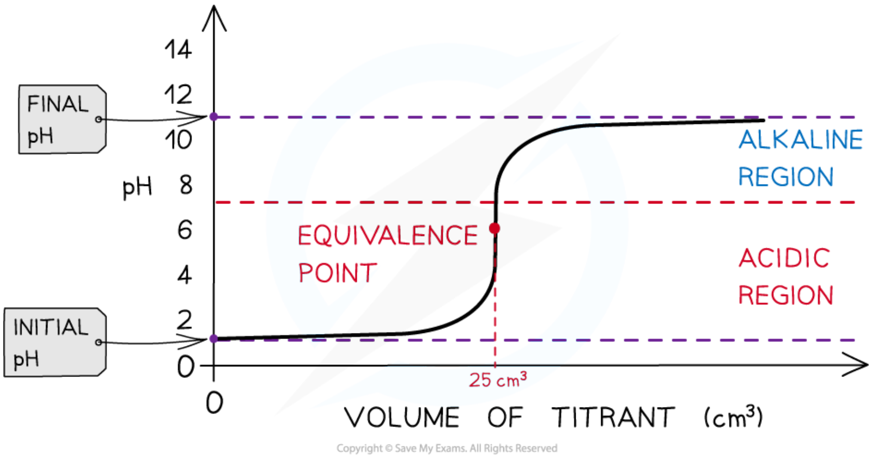


The diagram shows a pH titration curve of sodium hydroxide with hydrochloric acid

Strong acid + weak alkali pH titration curve

- Initially, there are **only H^+ ions present** in the solution from the dissociation of the strong acid (HCl) (initial pH about 1–2)
- As the volume of weak alkali (NH_3) added increases, the **pH of the analyte solution slightly increases too** as more and more H^+ ions react with the NH_3
- The change in pH is not that much until the volume added gets close to the equivalence point
- The **equivalence point is the point at which all H^+ ions have been neutralised by the NH_3** however the equivalence point is not neutral, but the solution is still acidic (pH about 5.5)
- This is because all H^+ have reacted with NH_3 to form NH_4^+ which is a relatively strong acid, causing the solution to be acidic
- As more of the NH_3 is added, the pH increases to above 7 but below that of a strong alkali as NH_3 is a weak alkali

pH titration curve for a strong acid + weak alkali



The diagram shows a pH titration curve of hydrochloric acid with ammonia

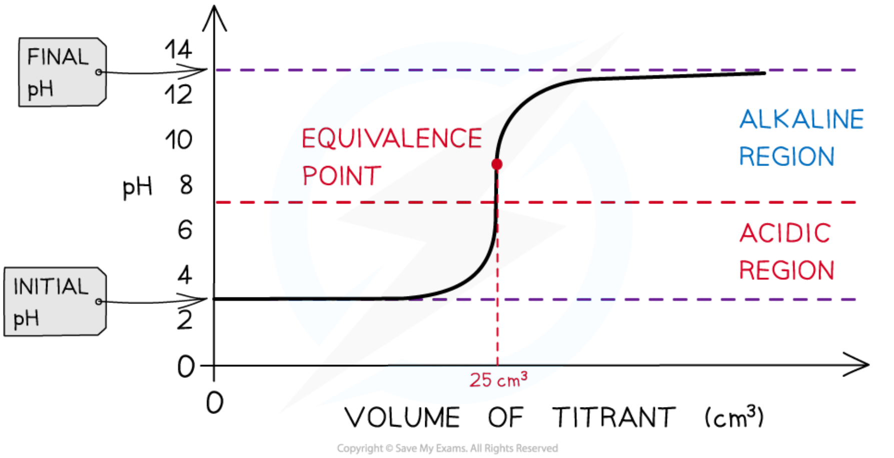
- The pH titration curve for strong acid added to a weak alkali has the same shape
- The initial and final pH are the other way around
- The **equivalence point is still about 5.5**

Weak acid + strong alkali pH titration curve

- Initially, there are **only H^+ ions present in the solution** from the dissociation of the weak acid (CH_3COOH, ethanoic acid) (initial pH about 2–3)
- As the volume of strong alkali (NaOH) added increases, the **pH of the ethanoic acid solution slightly increases too** as more and more H^+ ions react with the OH^- from the NaOH to form water
- The change in pH is not that much until the volume added gets close to the equivalence point
- The pH surges upwards very steeply
- The **equivalence point is the point at which all H^+ ions have been neutralised by the OH^- ions** however the equivalence point is not neutral, but the solution is slightly basic (pH about 9)
- This is because all H^+ in CH_3COOH have reacted with OH^- however, CH_3COO^- is a relatively strong base, causing the solution to be basic
- As more of the NaOH is added, the pH increases to about 13–14

pH titration curve for a weak acid + strong alkali

Your notes



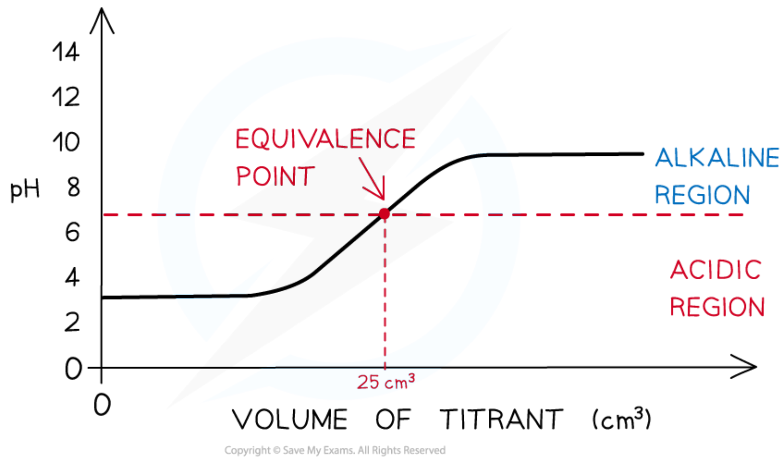
The diagram shows a pH titration curve of a weak acid with a strong base

- The pH titration curve for weak acid added to a strong alkali has the same shape
- The initial and final pH are the other way around
- The **equivalence point is still about 9**

Weak acid + weak alkali pH titration curve

- Initially, there are only H^+ ions present in the solution from the dissociation of the weak acid (CH_3COOH, ethanoic acid) (initial pH about 2–3)
- In these pH titration curves, there is no vertical region
- There is a **'point of inflexion'** at the equivalence point
- The curve does not provide much other information

pH titration curve for a weak acid + weak alkali



The diagram shows a pH titration curve of weak acid with weak alkali

Your notes

Your notes

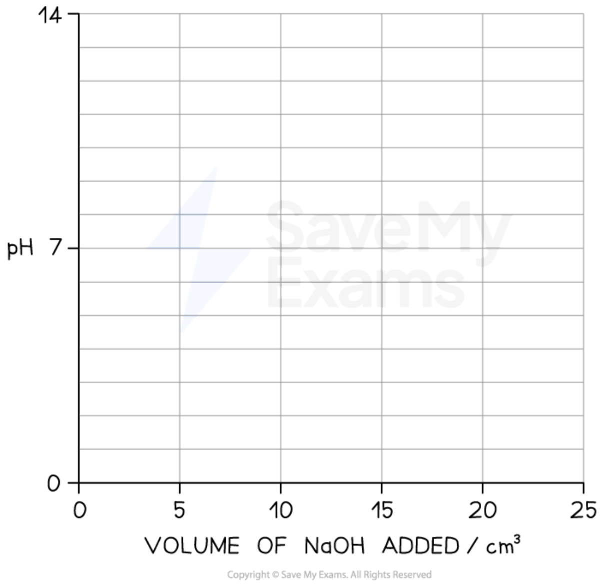
Examiner Tips and Tricks

You should be able to read and sketch pH titration curves of titrations where the titrant is an acid or an alkali.

Worked Example

A 10.0 cm^3 sample of 0.150 mol dm^{-3} aminoethanoic acid with a pH of 5.3 was titrated with 0.100 mol dm^{-3} NaOH. After 20.0 cm^3 of NaOH, an excess, had been added, the pH was found to be 12.5.

Using the following axes, sketch a graph showing how the pH changes during this titration.

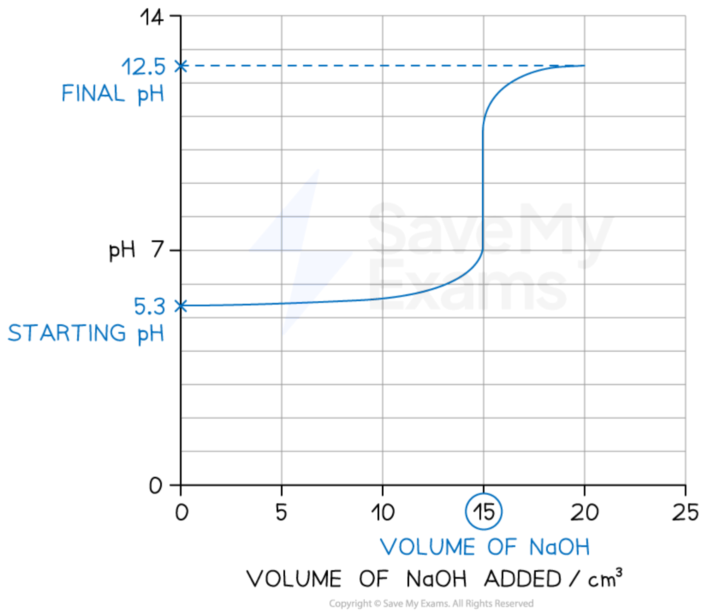


[3]

Answer

- The curve starts at pH 5.3
    - Mark on graph
- The volume of NaOH added to reach the vertical section of the graph = 15.0 cm^3
    - Vol acid x Concentration acid = Vol base x Concentration base
    - 10 x 0.150 = Vol base x 0.100
    - $\frac{(10 \times 0.150)}{0.100}$ = 15.0 cm^3
    - There is no mark for the height of the vertical section, but the equivalence point must be **above** pH 7 for a weak acid - strong base titration
- The curve finishes at pH = 12.5 at 20 cm^3.

- Make sure the graph does not go above pH 12.5
  - This is the maximum pH value given in the question
- Make sure that the volume does not exceed 20 cm^3
  - This is the maximum volume of base added given in the question



Your notes

Indicators used in Titration

Your notes

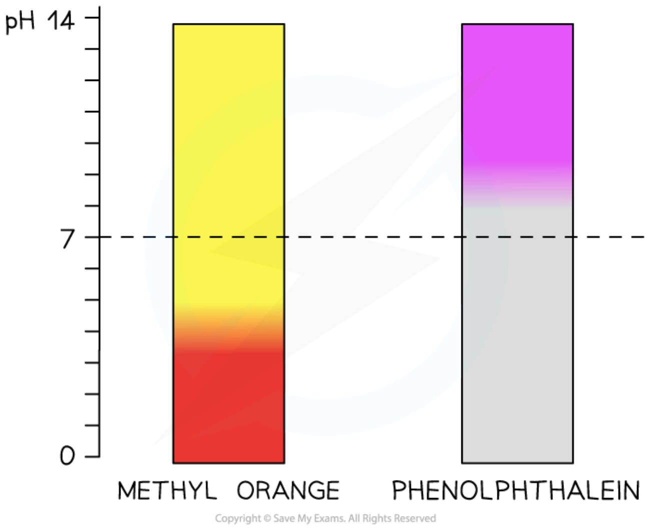
Indicators

- **Indicators** are substances that change colour when they are added to acidic or alkaline solutions
- When choosing the appropriate indicator, the pH of the equivalence point is very important
- The two most common indicators that are used in titrations are **methyl orange** and **phenolphthalein**

Indicator & pH range examples

- Both indicators change colour over a specific pH range
  - **Methyl orange** 3.1 - 4.4
  - **Phenolphthalein** 8.3 - 10.0

Diagram showing the colour changes for methyl orange and phenolphthalein



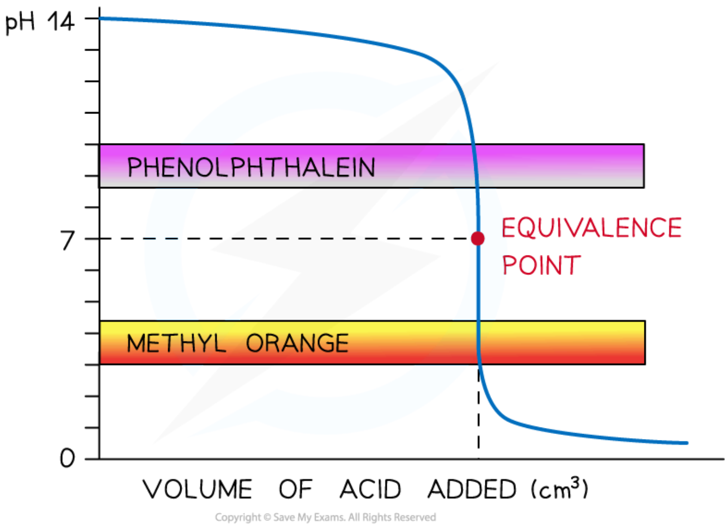
Methyl orange changes from red to yellow over a pH range of 3.1 – 4.4, while phenolphthalein changes from colourless to pink over a pH range of 8.3 – 10.0

Choosing indicators for titrations

Strong acid and strong alkali

- The colour change for **both indicators** takes place at a pH range that falls within the vertical region of the curve
- Therefore, either indicator can be used

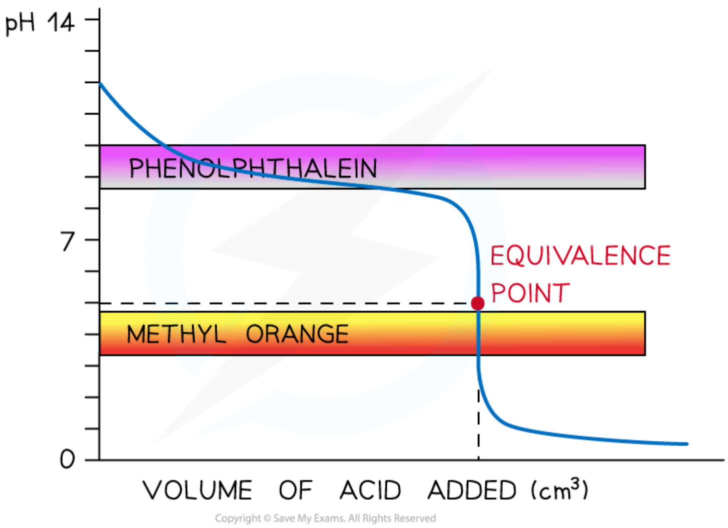
Methyl orange and phenolphthalein in a strong acid + strong alkali titration



Both indicators can be used to determine the endpoint of the titration of a strong acid and strong alkali

- **Strong acid and weak alkali**
  - Only **methyl orange** will change colour at a pH close to the equivalence point and within the vertical region of the curve

Methyl orange and phenolphthalein in a strong acid + weak alkali titration



Only methyl orange can be used to determine the endpoint of the titration of a strong acid and weak alkali

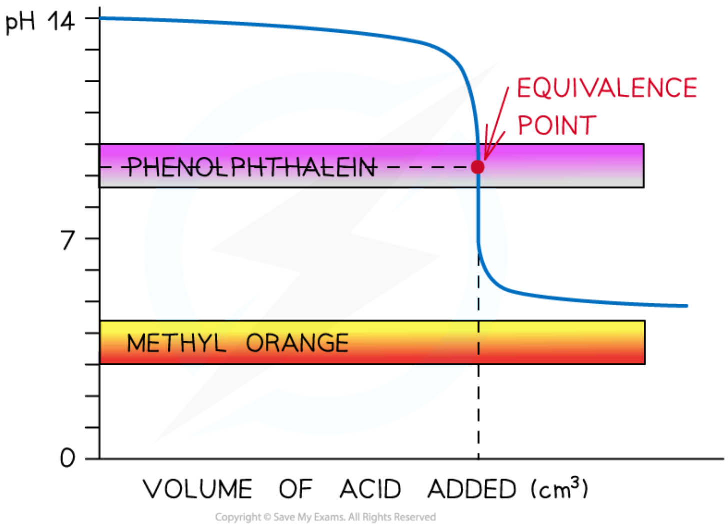
- **Weak acid and strong alkali**

Your notes

- Now, only **phenolphthalein** will change colour at a pH close to the equivalence point and within the vertical region of the curve
- The pH range at which methyl orange changes colour falls below the curve

Your notes

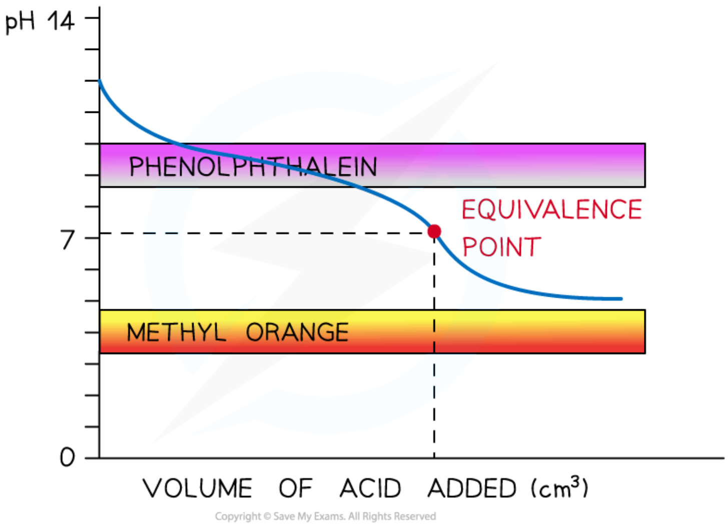
Methyl orange and phenolphthalein in a weak acid + strong alkali titration



Only phenolphthalein can be used to determine the endpoint of the titration of a weak acid and strong alkali

- **Weak acid and weak alkali**
  - Neither indicator is useful, and a different method should be considered

Methyl orange and phenolphthalein in a weak acid + weak alkali titration



Neither indicator can be used to determine the endpoint of the titration of a weak acid and weak alkali