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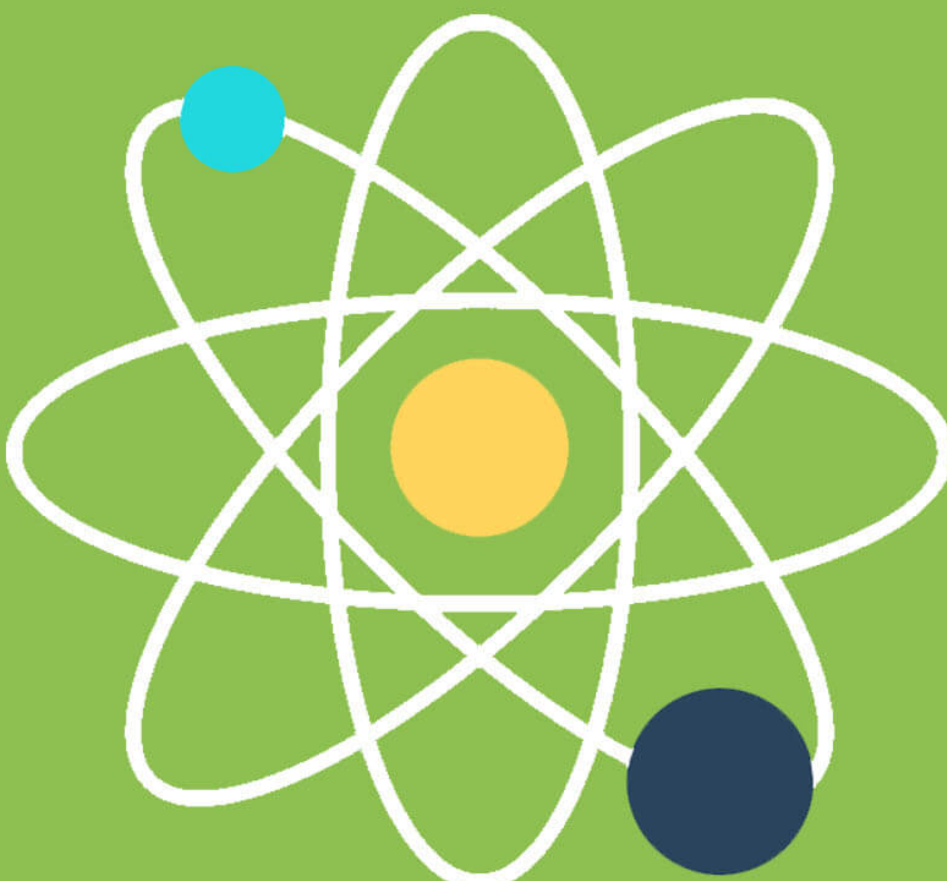
8.2 More About Acids



IB Chemistry - Revision Notes

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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 $\mathbf{H^+}$ 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

- 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

8.2.2 pH & [H+]

pH & [H+]

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

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

  - 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

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**

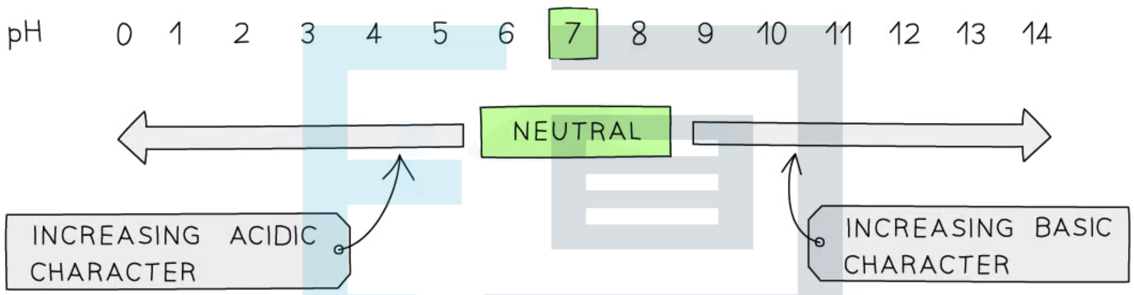
Exam Tip

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

8.2.3 Interpreting pH

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



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 $\mathbf{H^+}$ is always **greater** than the concentration of $\mathbf{OH^-}$ 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 $\mathbf{OH^-}$ is always **greater** than the concentration of $\mathbf{H^+}$ 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:

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

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

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

$$= 7$$

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

8.2.4 The Ionic Product of Water

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

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**

8.2.5 Acid-Base Calculations

Acid-Base Calculations

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

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

- 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 x 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 x 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 x 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 x 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:

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

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

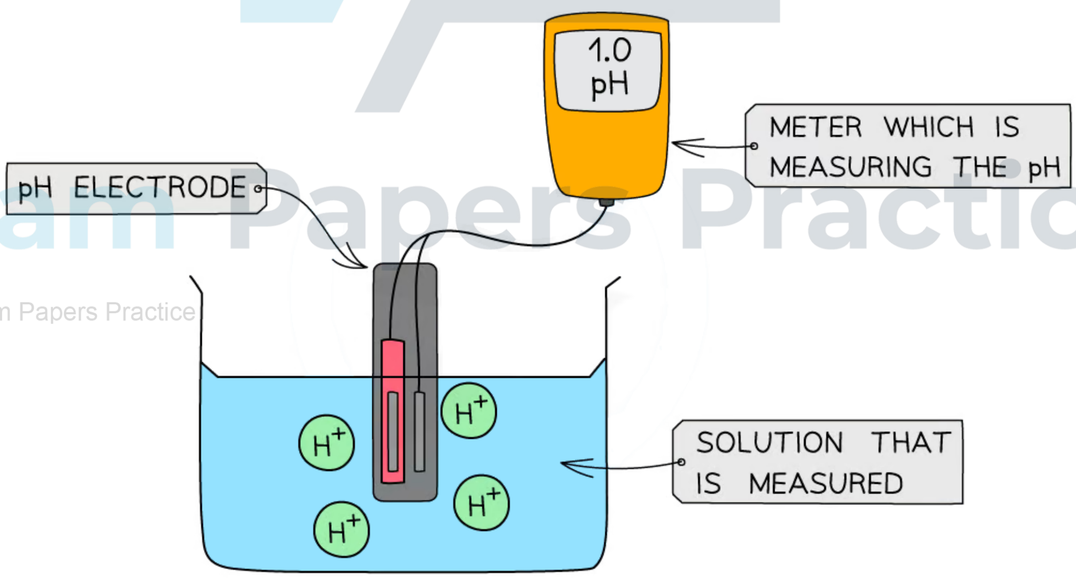
$$[H^+] = \frac{1 \times 10^{-14}\ mol^2 dm^{-6}}{0.357\ mol\ dm^{-3}} = 2.80 \times 10^{-14} \text{ mol dm}^{-3}$$

pH = - log (2.80 x 10^{-14}) = 13.55

8.2.6 pH Meters & Universal Indicator

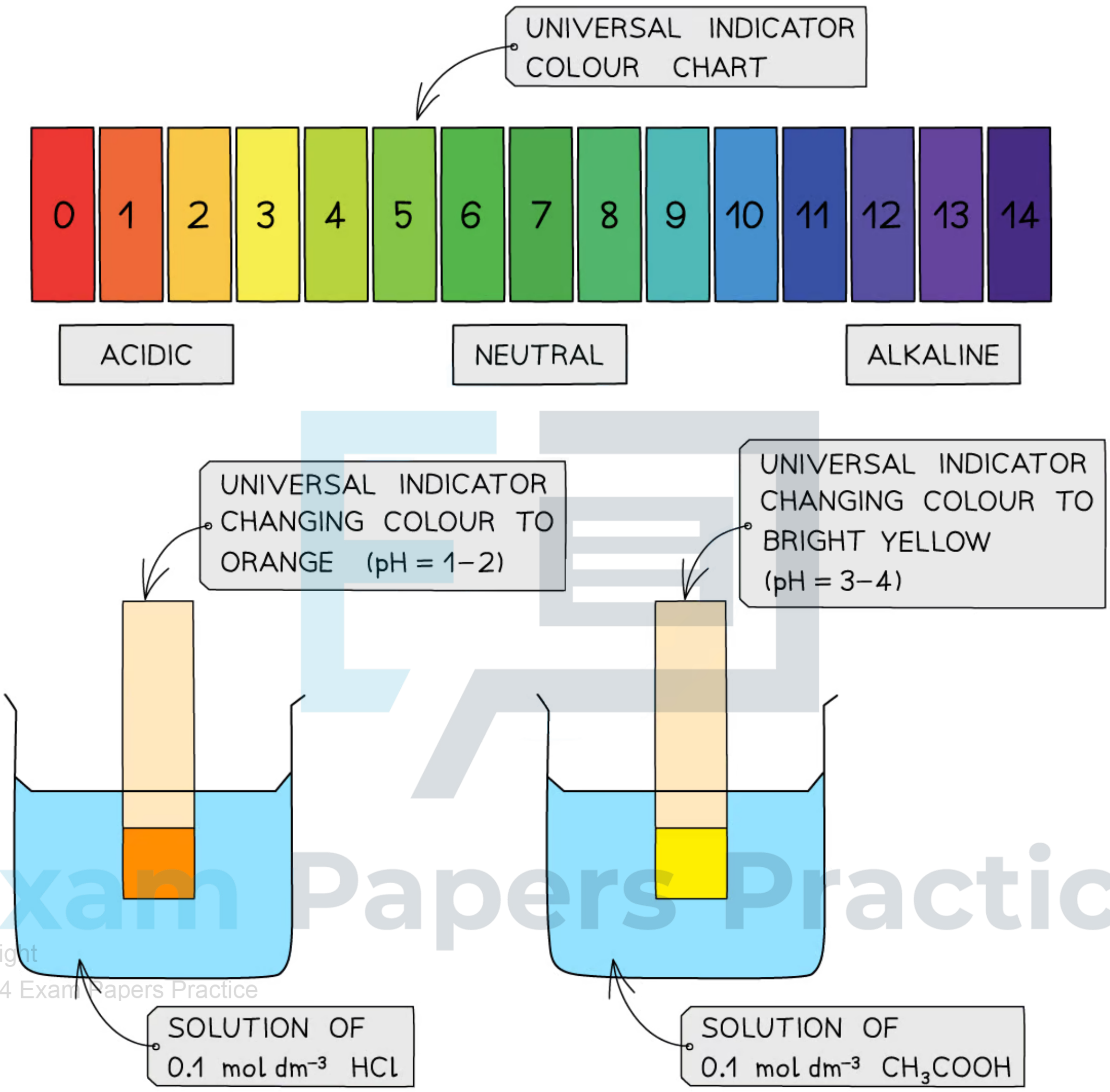
pH Meters & Universal Indicator

- 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



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



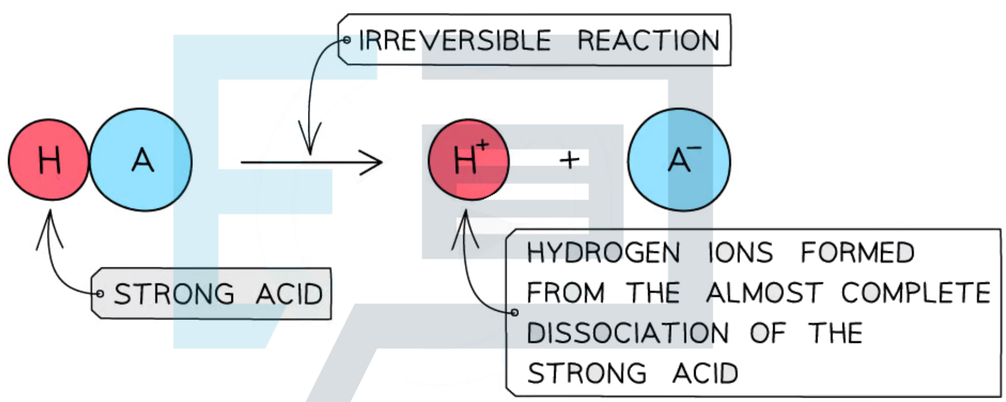
The diagram shows the change in colour of the universal indicator paper when dipped in a strong (HCl) and weak (CH_3COOH) acid. The colour chart is used to read off the corresponding pH values which are between 1–2 for HCl and 3–4 for CH_3COOH

8.2.7 Strong & Weak Acids & Bases

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



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

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

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

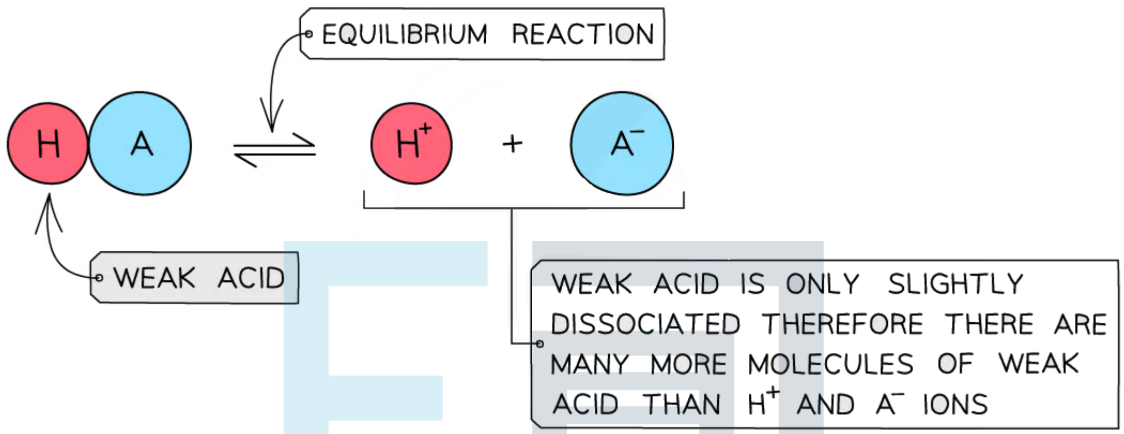
pH is the negative log of the concentration of H^+/H_3O^+ ions and can be calculated if the concentration of the strong acid is known using the stoichiometry of the reaction

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



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

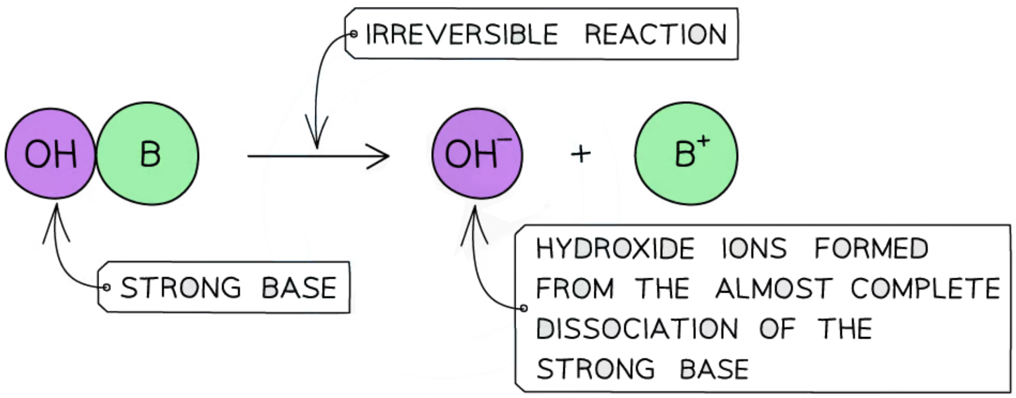
- The solution is **less acidic** due to the lower concentration of H^+/H_3O^+ ions
- Finding the pH of a weak acid requires using the acid dissociation constant, K_a but this not required at Standard Level, but only at Higher Level and is covered in Topic 18

Acid & Equilibrium Position Table

	Strong Acid	Weak Acid
Position of Equilibrium	Right	Left
Dissociation	Completely ($\rightarrow$)	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

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

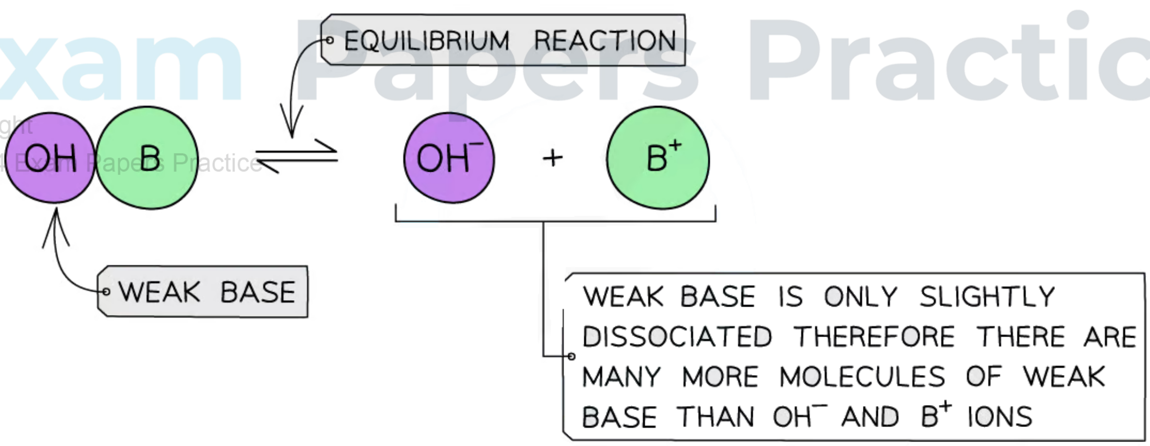


The diagram shows the complete dissociation of a strong base in aqueous solution

- 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

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

Conjugate Pairs & Acid-Base Strength

- 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.**

8.2.8 Comparing Strong & Weak Acids

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:

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

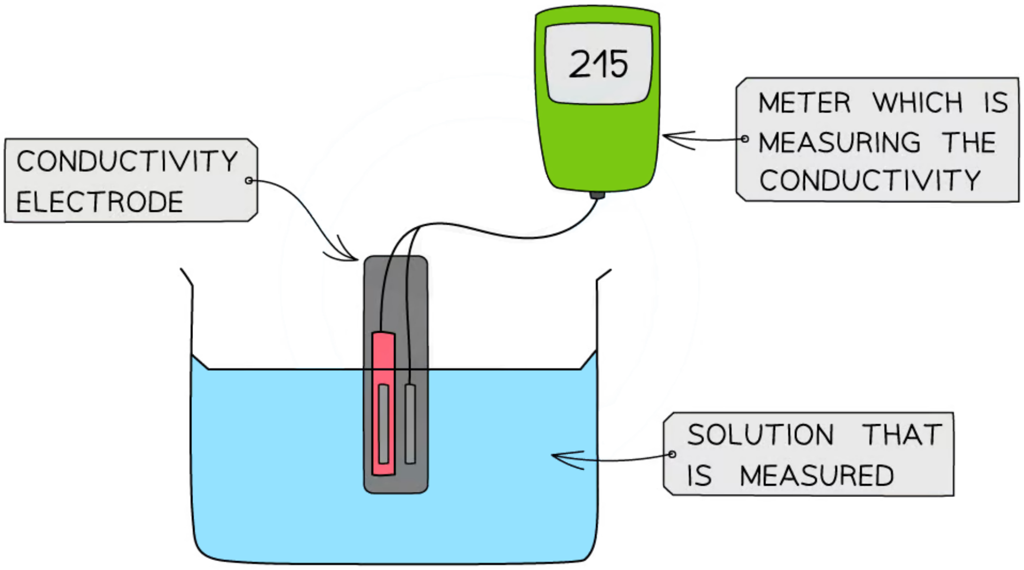
- The **stronger** the acid, the **greater** the **concentration of H^+** and therefore the **lower the pH**

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

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

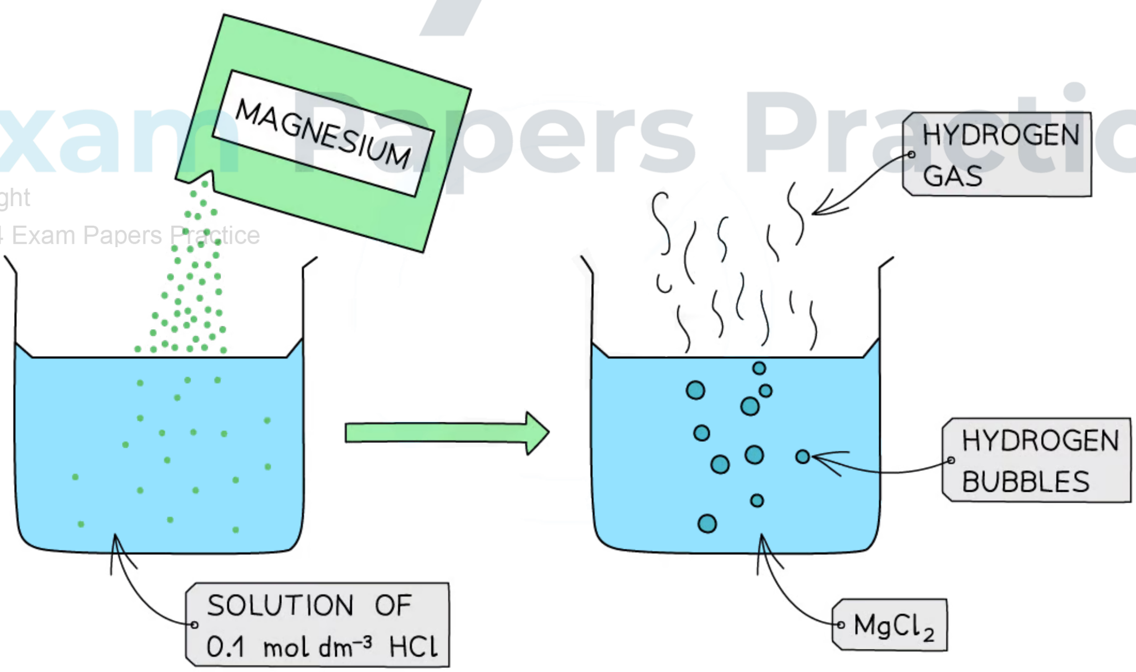


The diagram shows a digital conductivity meter that measures the electrical conductivity of a solution using an electrode

Reactivity

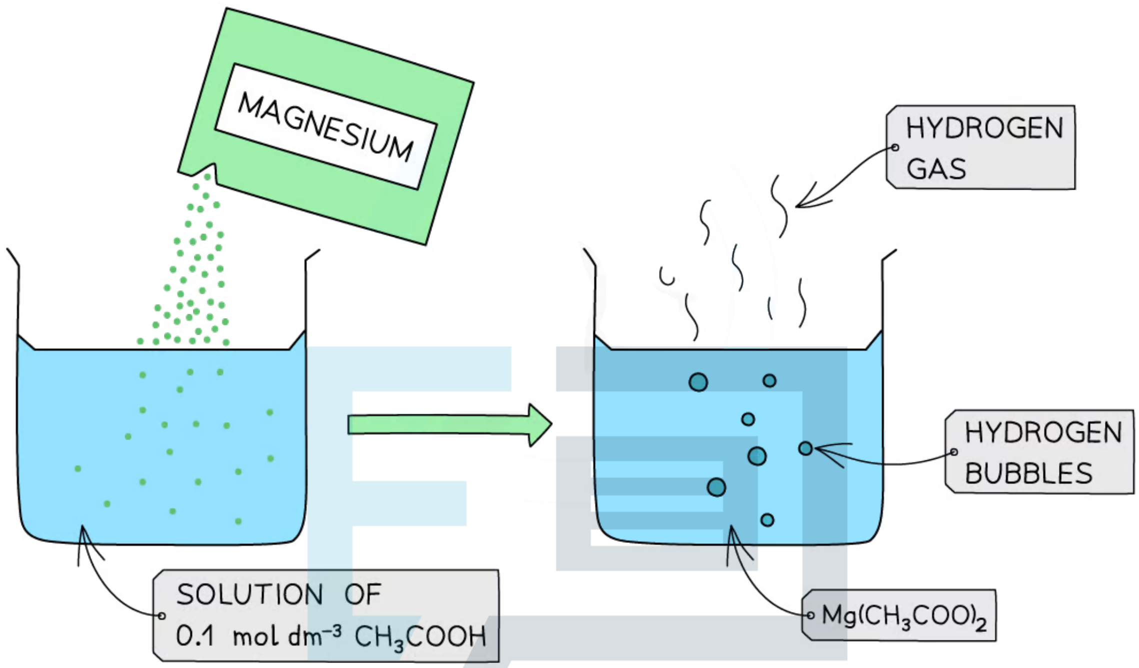
- Strong and weak acids of the **same concentrations** react differently with reactive metals
- This is because the concentration of H^+ is greater in strong acids compared to weak acids
- The greater H^+ concentration means that more H_2 gas is produced in a shorter time

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



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

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



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.