DP IB Chemistry: SL

Your notes

8.2 More About Acids

Contents

Your notes

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

- 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

Examiner Tip

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

Your notes

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

Your notes

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**

Examiner Tip

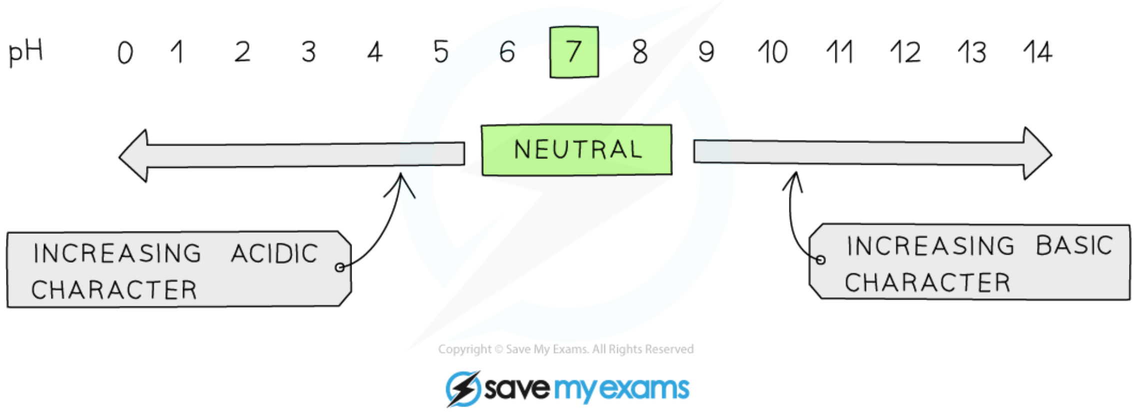
Make sure you know how to use the antilog (base 10) feature on your calculator. On most calculators it is the 10^x button, but on other models it could be LOG^{-1}, ALOG or even a two-button sequence such as INV + LOG

8.2.3 Interpreting pH

Your notes

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 **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

Your notes

- 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:

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

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

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

$$= 7$$

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

Your notes

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

Your notes

$[H^+]$	$[OH^-]$	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 = 1.0×10^{-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

Your notes

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:

Your notes

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

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

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

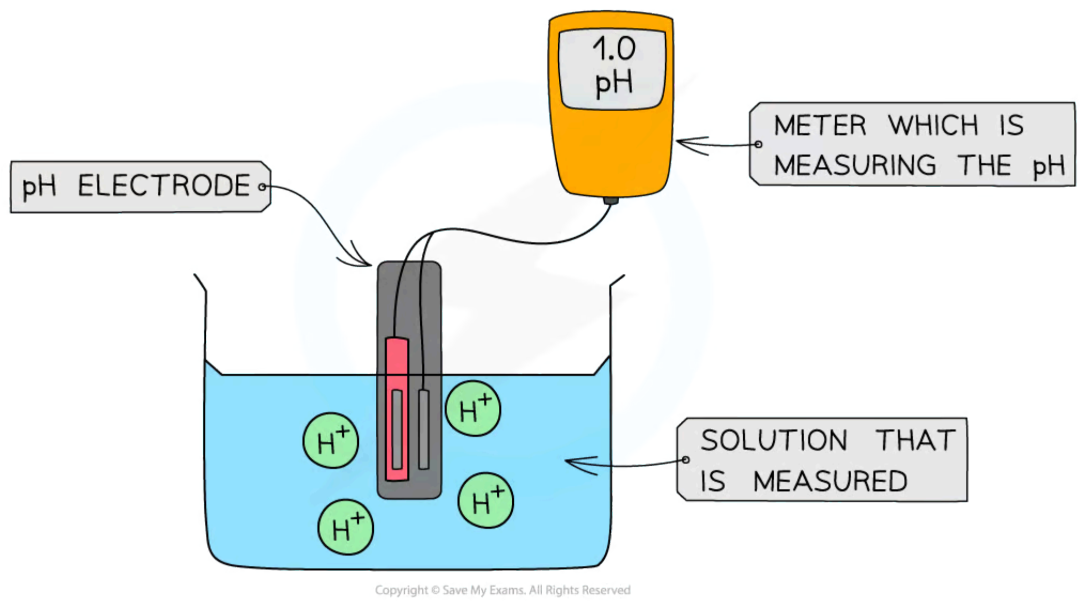
pH = – log (2.80 x 10^{-14}) = 13.55

8.2.6 pH Meters & Universal Indicator

Your notes

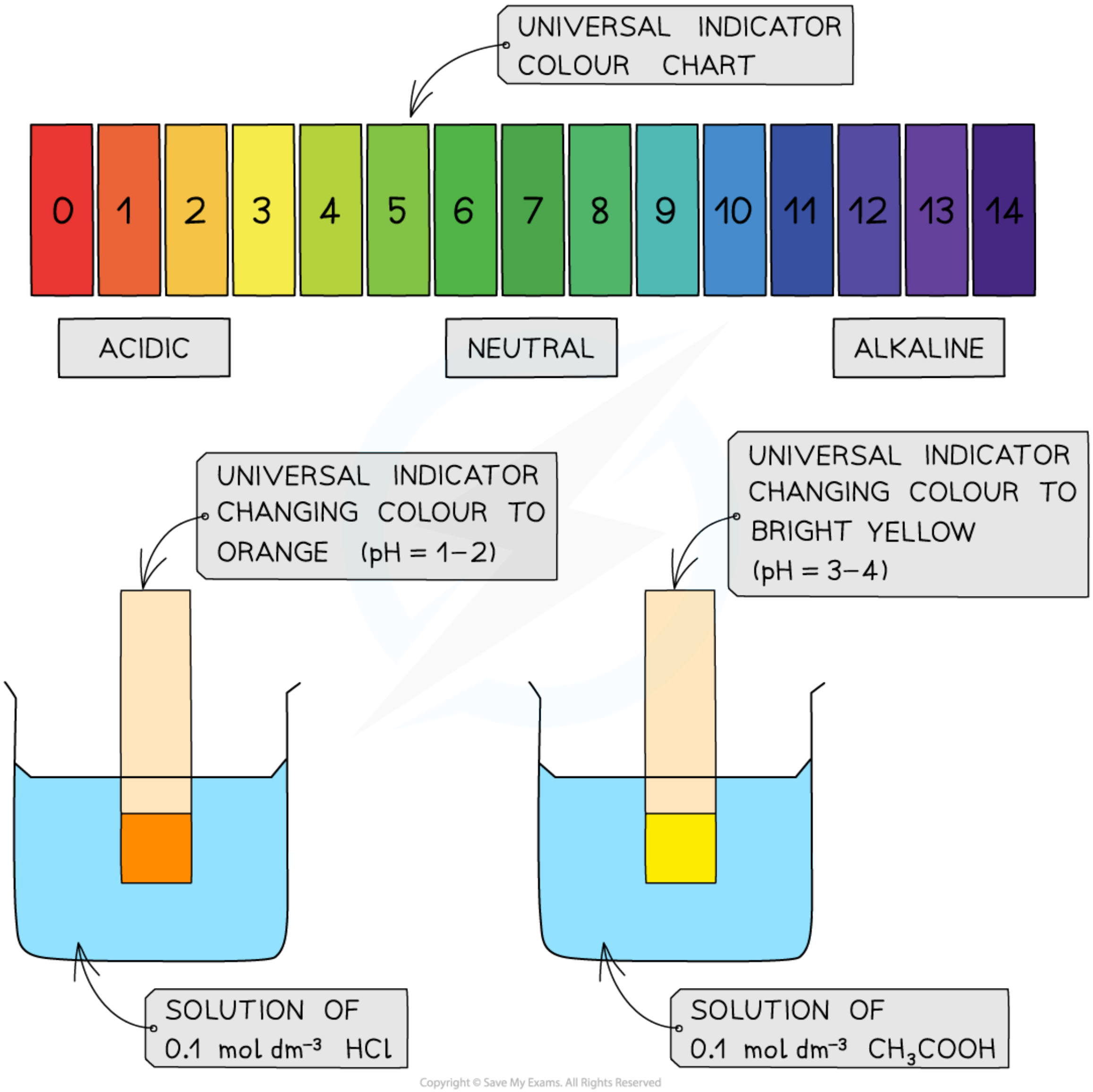
pH Meters & Universal Indicator

- 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

Your notes

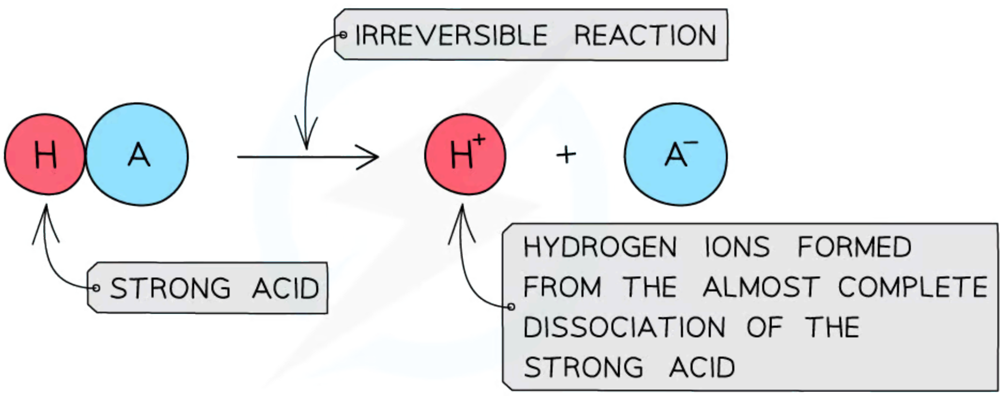
Your notes

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

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

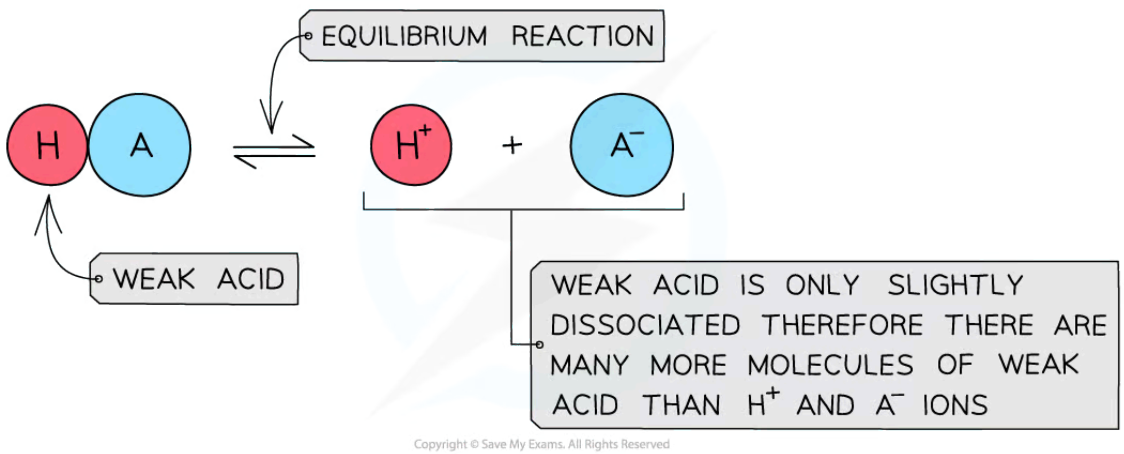
$$[H^+(aq)] = \text{CONCENTRATION OF } H^+/H_3O^+ \text{ IONS}$$

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

Your notes

- 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

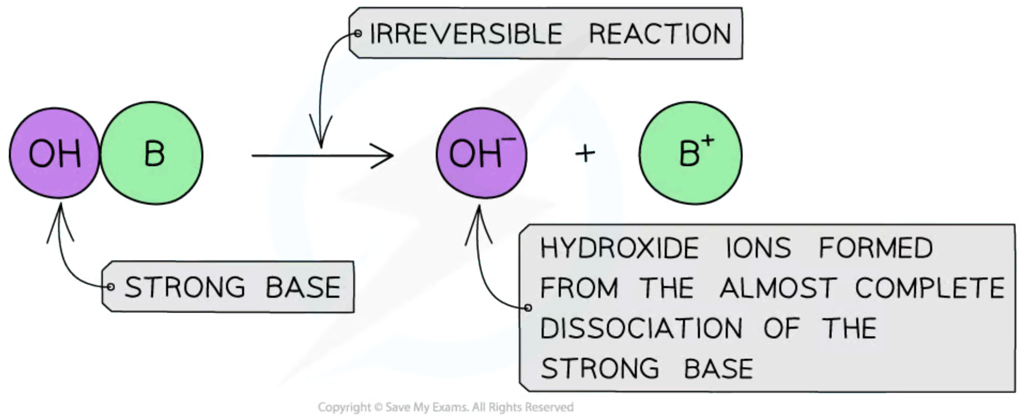
Your notes

	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

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

Your notes

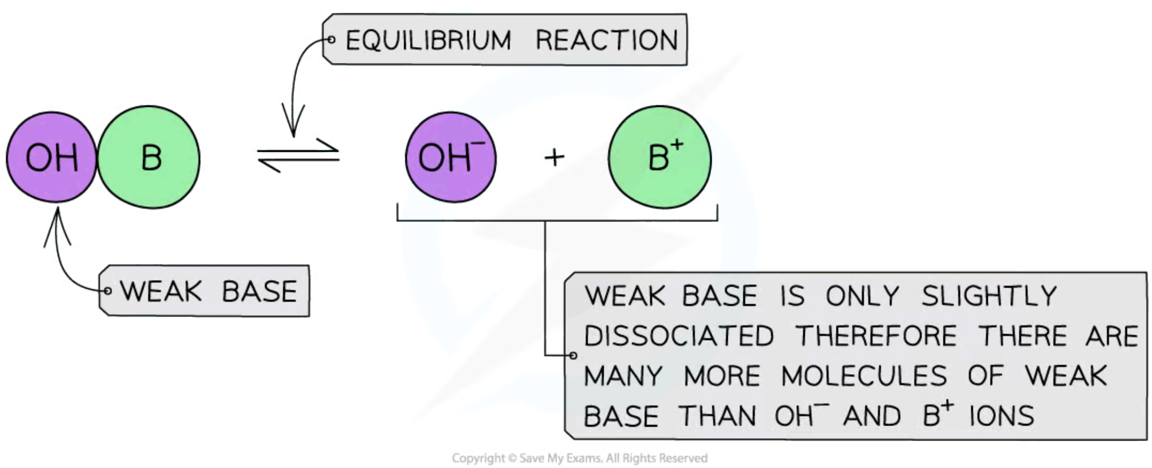


The diagram shows the complete dissociation of a strong base in aqueous solution

- 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

Your notes

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

Your notes

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**

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: **$H_2SO_4 \rightarrow H^+ + SO_4^-$**$HSO_4^-$ 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

Your notes

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**

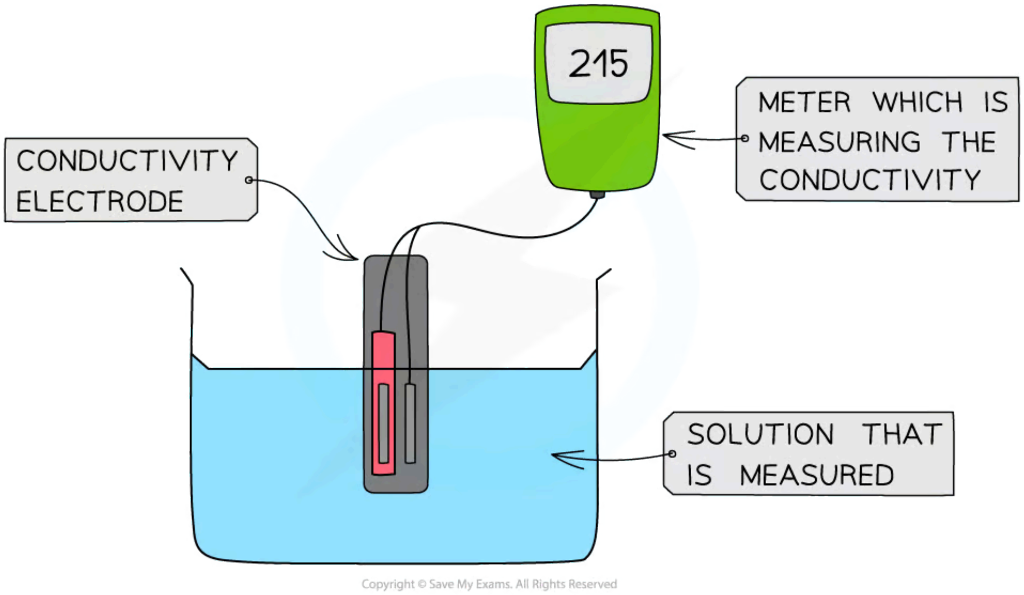
pH value of a Strong Acid & Weak Acid Table

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

Your notes

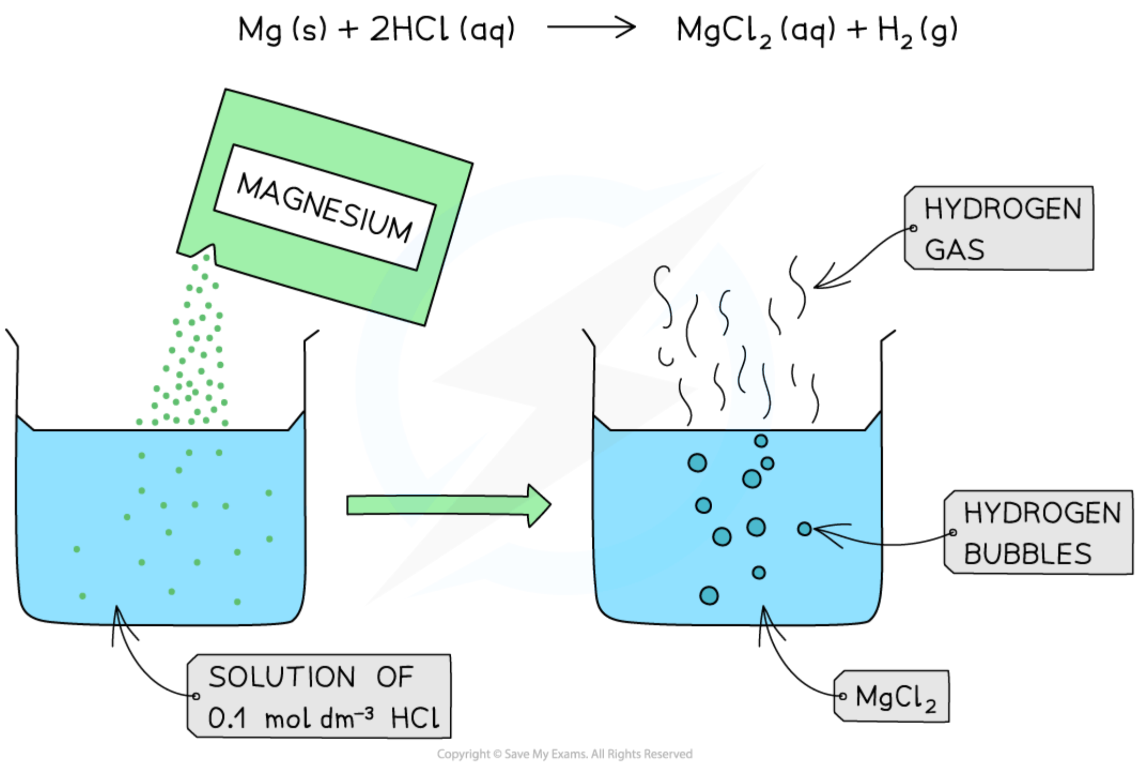


The diagram shows a digital conductivity meter that measures the electrical conductivity of a solution using an electrode

Reactivity

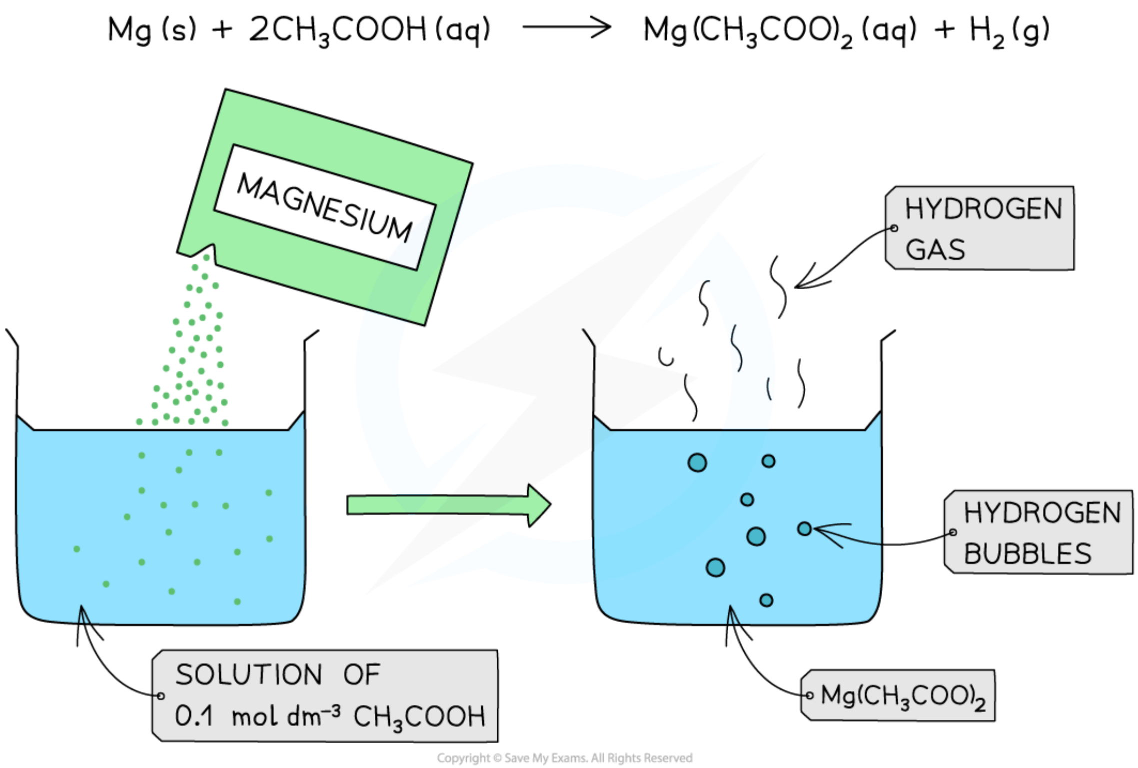
- Strong and weak acids of the **same concentrations** react differently with reactive metals
- This is because the concentration of H^+ is greater in strong acids compared to weak acids
- The greater H^+ concentration means that more H_2 gas is produced in a shorter time

Your notes



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

Your notes



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, producing a greater concentration of H^+ ions and therefore showing lower pH values, greater electrical conductivity and more vigorous reactions with reactive metals.