

# DP IB Chemistry: HL



## 8.2 More About Acids

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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  $\text{H}^+$  ions in the solution
- **Indicators** are weak acids and bases where the conjugate bases and acids have a different colour
- Many **acid-base indicators** are derived from plants, such as litmus

Common Indicators Table

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

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

#### Examiner Tip

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



Your notes

## 8.2.2 pH & [H<sup>+</sup>]

### pH & [H<sup>+</sup>]

- The acidity of an aqueous solution depends on the number of H<sup>+</sup> (H<sub>3</sub>O<sup>+</sup>) ions in solution
- The pH is defined as:

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

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- where [H<sup>+</sup>] is the concentration of H<sup>+</sup> in mol dm<sup>-3</sup>
- 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<sup>+</sup>] Table

[H <sup>+</sup> ]	Scientific notation	pH
1.0	10 <sup>0</sup>	0
0.1	10 <sup>-1</sup>	1
0.01	10 <sup>-2</sup>	2
0.001	10 <sup>-3</sup>	3
0.0001	10 <sup>-4</sup>	4
–	10 <sup>-x</sup>	x

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Your notes

### Worked example

10.0 cm<sup>3</sup> of an aqueous solution of nitric acid of pH = 1.0 is mixed with 990.0 cm<sup>3</sup> 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<sup>3</sup> so the concentration of H<sup>+</sup> has been **reduced** by a factor of 100 or 10<sup>-2</sup>, 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<sup>x</sup> button, but on other models it could be LOG<sup>-1</sup>, ALOG or even a two-button sequence such as INV + LOG

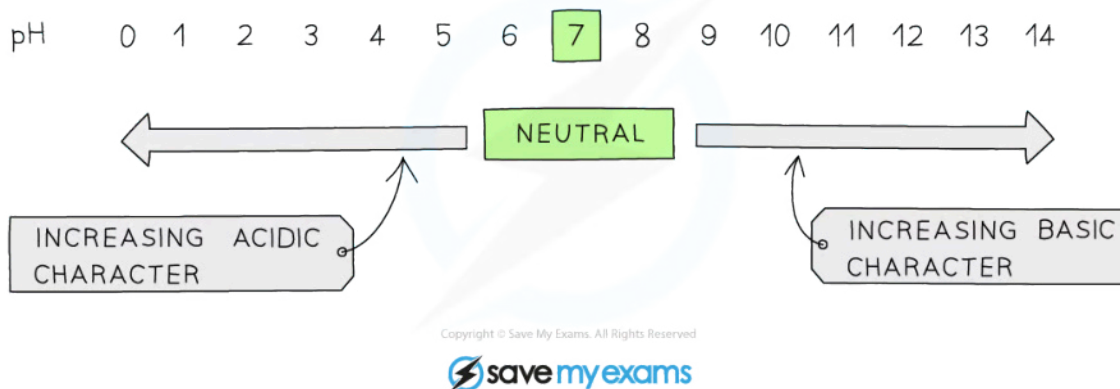


Your notes

## 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  $\text{H}^+$  than  $\text{OH}^-$  ions
- Since the concentration of  $\text{H}^+$  is always **greater** than the concentration of  $\text{OH}^-$  ions,  $[\text{H}^+]$  is always **greater** than  $10^{-7} \text{ mol dm}^{-3}$
- Using the pH formula, this means that the **pH of acidic solutions** is always **below 7**
- The higher the  $[\text{H}^+]$  of the acid, the lower the pH

### pH of bases

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

### The pH of water

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

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

$$[\text{H}^+ (\text{aq})] = \text{CONCENTRATION OF } \text{H}^+/\text{H}_3\text{O}^+ \text{ IONS}$$

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$$\begin{aligned} \text{pH} &= -\log(10^{-7}) \\ &= 7 \end{aligned}$$

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



Your notes



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



- The equilibrium constant for this reaction is:

$$K_c = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

$$K_c \times [\text{H}_2\text{O}] = [\text{H}^+][\text{OH}^-]$$

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

$$K_w = [\text{H}^+][\text{OH}^-]$$

Where  $K_w$  (ionic product of water) =  $K_c \times [\text{H}_2\text{O}]$

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

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

**[H<sup>+</sup>] & [OH<sup>-</sup>] Table**



Your notes

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

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### Worked example

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

- 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  $(1.0 \times 10^{-14}) \div (1.0 \times 10^{-3}) = 1.0 \times 10^{-11} \text{ mol dm}^{-3}$
- So the **pH = 11**





Your notes

## 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

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

#### Answers:

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

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

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

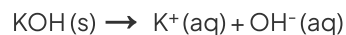
The final pH is 5, which corresponds to  $1 \times 10^{-5} \text{ mol dm}^{-3}$

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

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

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

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



The concentration of KOH is

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

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

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

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



Your notes

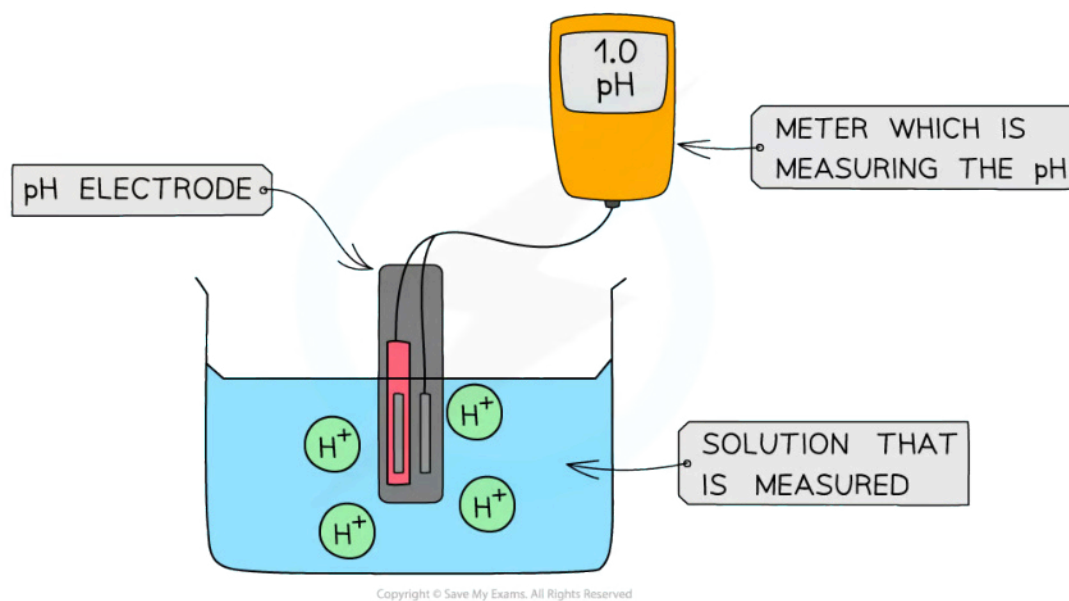


Your notes

## 8.2.6 pH Meters & Universal Indicator

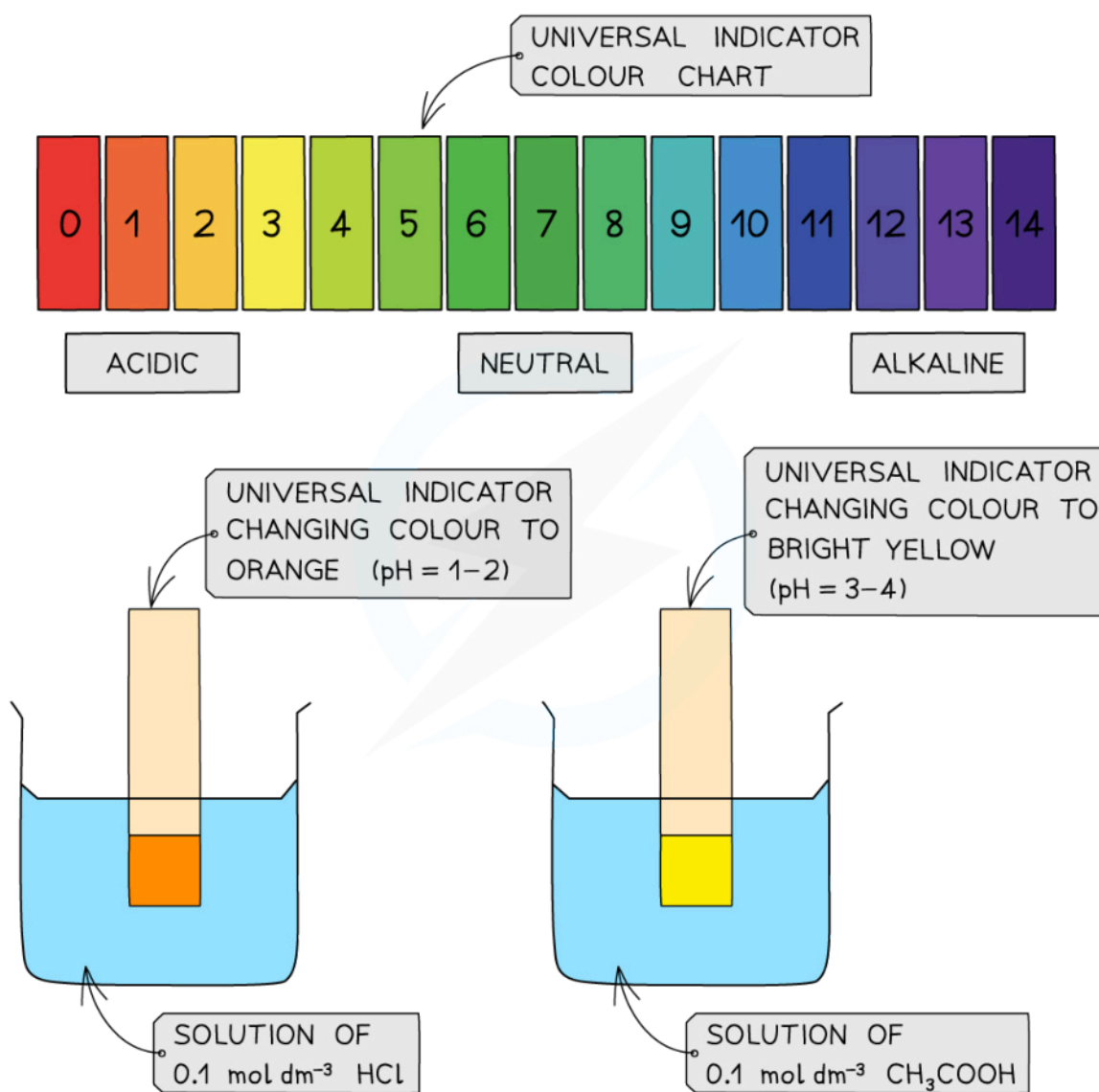
### pH Meters & Universal Indicator

- 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



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The diagram shows the change in colour of the universal indicator paper when dipped in a strong (HCl) and weak (CH<sub>3</sub>COOH) 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<sub>3</sub>COOH



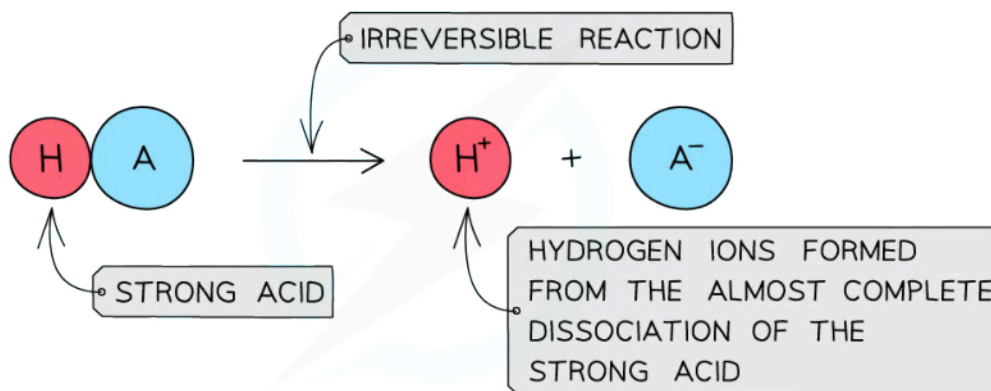
Your notes

## 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<sub>3</sub> (nitric acid) and H<sub>2</sub>SO<sub>4</sub> (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<sup>+</sup>/H<sub>3</sub>O<sup>+</sup> ions
- Since the **pH** depends on the concentration of H<sup>+</sup>/H<sub>3</sub>O<sup>+</sup> ions, the pH can be calculated if the concentration of the strong acid is known

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

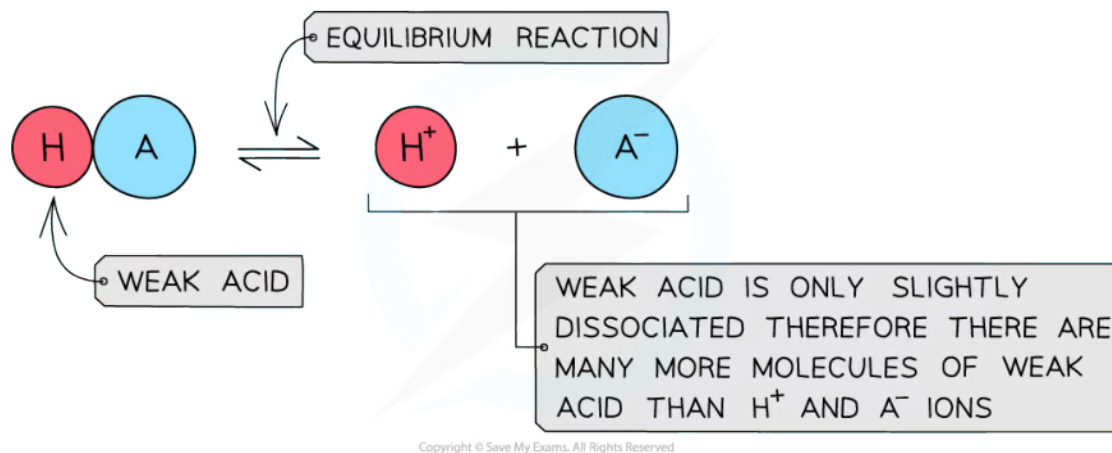
$$[\text{H}^+ (\text{aq})] = \text{CONCENTRATION OF H}^+/\text{H}_3\text{O}^+ \text{ IONS}$$

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*pH is the negative log of the concentration of H<sup>+</sup>/H<sub>3</sub>O<sup>+</sup> 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<sub>2</sub>S (hydrogen sulfide) and H<sub>2</sub>CO<sub>3</sub> (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<sup>+</sup>/H<sub>3</sub>O<sup>+</sup> ions
- Finding the pH of a weak acid requires using the acid dissociation constant, K<sub>a</sub> but this not required at Standard Level, but only at Higher Level and is covered in Topic 18

### Acid & Equilibrium Position Table



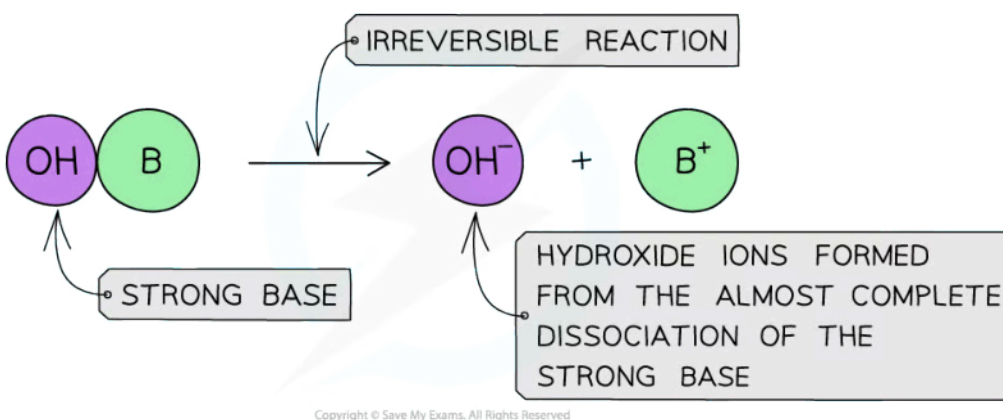
Your notes

	Strong Acid	Weak Acid
Position of Equilibrium	Right	Left
Dissociation	Completely ( $\rightarrow$ )	Partially ( $\rightleftharpoons$ )
$H^+$ concentration	High	Low
pH	Use [strong acid] for $[H^+]$	Use $K_a$ to find $[H^+]$
Examples	HCl HNO <sub>3</sub> H <sub>2</sub> SO <sub>4</sub> (first ionisation)	Organic acids (ethanoic acid) HCN H <sub>2</sub> S H <sub>2</sub> CO <sub>3</sub>

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### 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

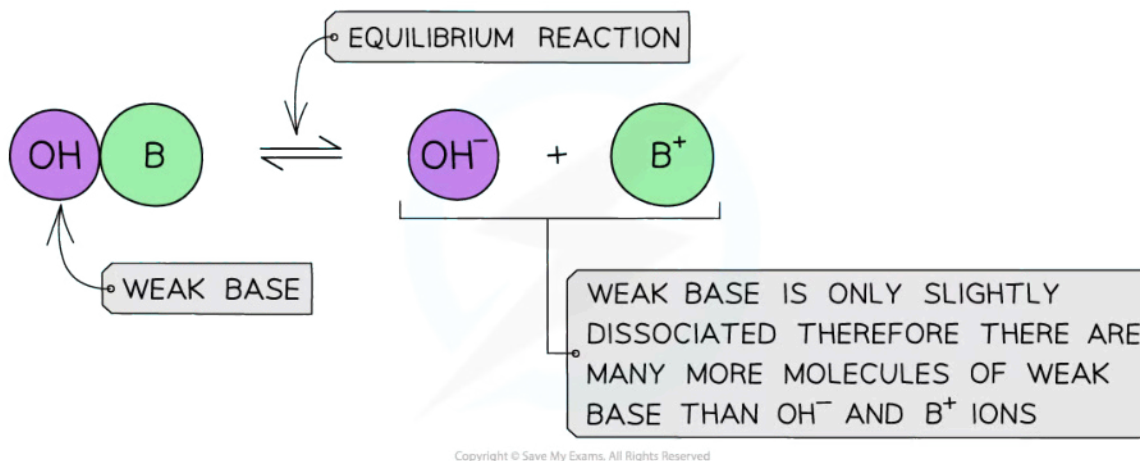


*The diagram shows the complete dissociation of a strong base in aqueous solution*

- The solution formed is highly basic due to the high concentration of the  $\text{OH}^-$  ions

### Weak bases

- A **weak base** is a base that **partially** (or incompletely) **dissociates** in aqueous solutions
  - $\text{NH}_3$  (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  $\text{OH}^-$  ions



Base & Equilibrium Position Table



Your notes

	Strong Base	Weak Base
Position of Equilibrium	Right	Left
Dissociation	Completely ( $\rightarrow$ )	Partially ( $\rightleftharpoons$ )
$\text{OH}^-$ concentration	High	Low
Examples	Group 1 metal hydroxides	$\text{NH}_3$ Amines Some transition metal hydroxides

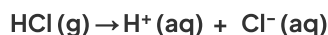
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## Conjugate Pairs & Acid-Base Strength



Your notes

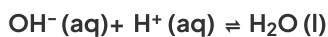
- The conjugate base of HCl is the chloride ion, Cl<sup>-</sup>, but since the reverse reaction is virtually non-existent the chloride ion must be a very weak conjugate base



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:



- 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<sub>3</sub>O<sup>+</sup> or as H<sup>+</sup> however, if H<sub>3</sub>O<sup>+</sup> is used, H<sub>2</sub>O should be included in the chemical equation: **HCl(g) → H<sup>+</sup>(aq) + Cl<sup>-</sup>(aq)** OR **HCl(g) + H<sub>2</sub>O(l) → H<sub>3</sub>O<sup>+</sup>(aq) + Cl<sup>-</sup>(aq)** Some acids contain two replaceable protons (called '**dibasic**') – for example, H<sub>2</sub>SO<sub>4</sub> (sulfuric acid) has two ionisations: H<sub>2</sub>SO<sub>4</sub> acts as a strong acid: **H<sub>2</sub>SO<sub>4</sub> → H<sup>+</sup> + SO<sub>4</sub><sup>2-</sup>** HSO<sub>4</sub><sup>-</sup> acts as a weak acid: **HSO<sub>4</sub><sup>-</sup> ⇌ H<sup>+</sup> + SO<sub>4</sub><sup>2-</sup>** The second ionisation is only partial which is why the concentration of 1 mol dm<sup>-3</sup> sulfuric acid is not 2 mol dm<sup>-3</sup> in H<sup>+</sup> 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**.



Your notes

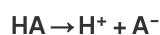
## 8.2.8 Comparing Strong & Weak Acids

### 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:



- 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 $0.1 \text{ mol dm}^{-3}$ solution
HCl (strong)	1
$CH_3COOH$ (weak)	2.9

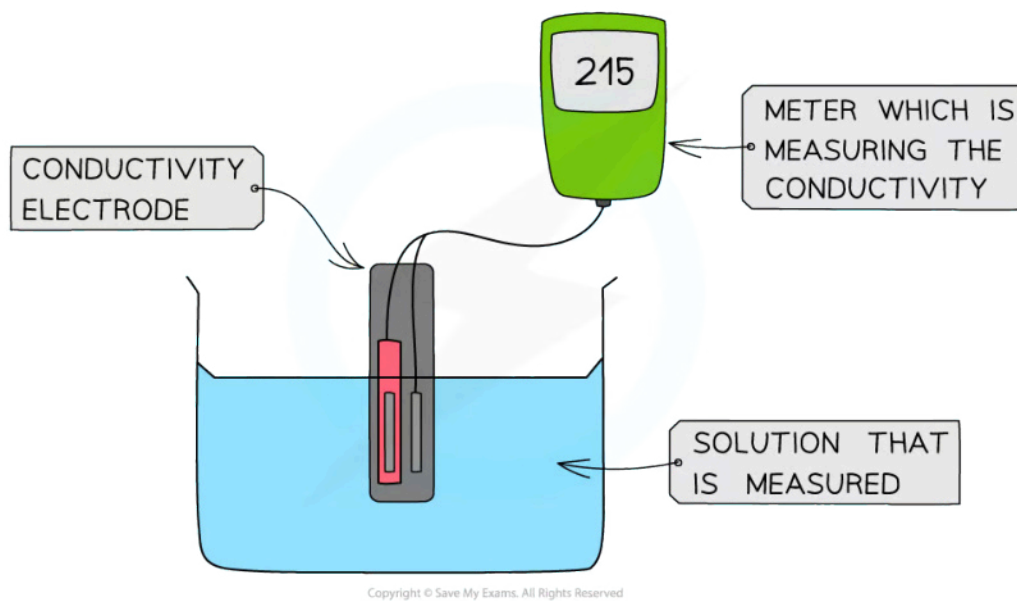
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#### Electrical conductivity

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



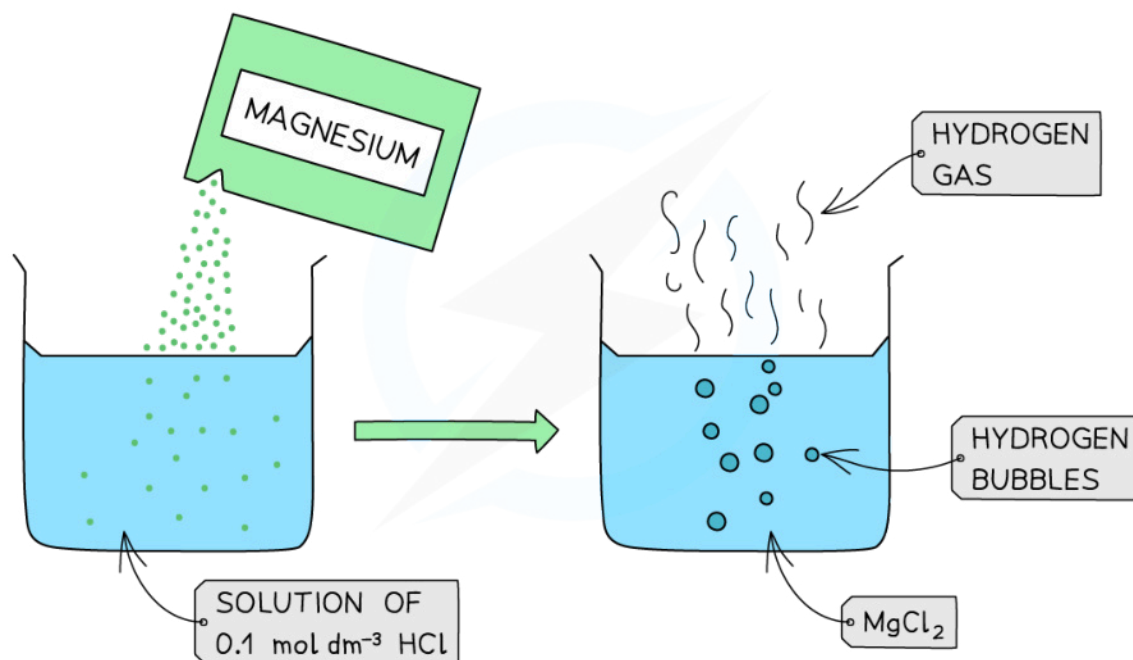
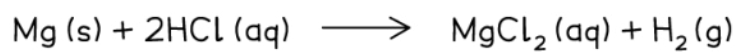
Your notes



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

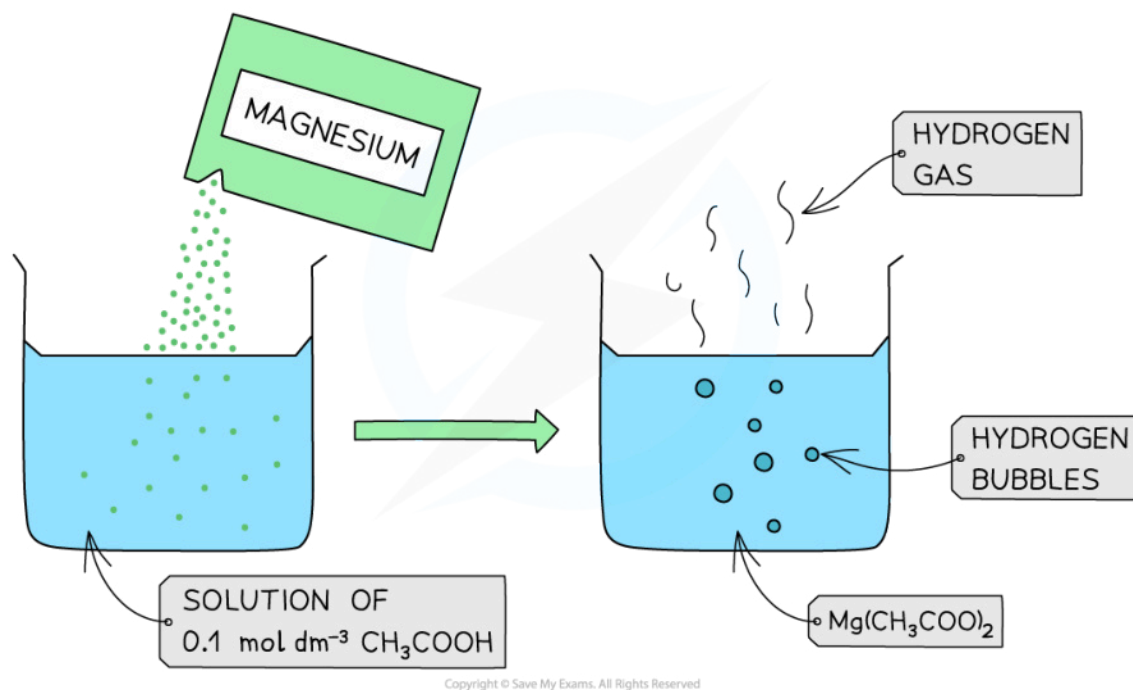
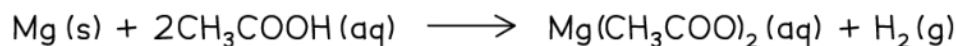
## Reactivity

- 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



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



The diagram shows the reaction of 0.1 mol dm<sup>-3</sup> of a weak acid (CH<sub>3</sub>COOH) with Mg. The reaction produces fewer bubbles of hydrogen gas due to the lower concentration of H<sup>+</sup> 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<sup>+</sup> ions. Stronger acids dissociate more, producing a greater concentration of H<sup>+</sup> ions and therefore showing lower pH values, greater electrical conductivity and more vigorous reactions with reactive metals.