

DP IB Chemistry: SL


Your notes

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 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⁺]

pH & [H⁺]

- The acidity of an aqueous solution depends on the number of H⁺ (H₃O⁺) ions in solution
- The pH is defined as:

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

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- 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

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Worked example

10.0 cm³ of an aqueous solution of nitric acid of pH = 1.0 is mixed with 990.0 cm³ 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 cm³ 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

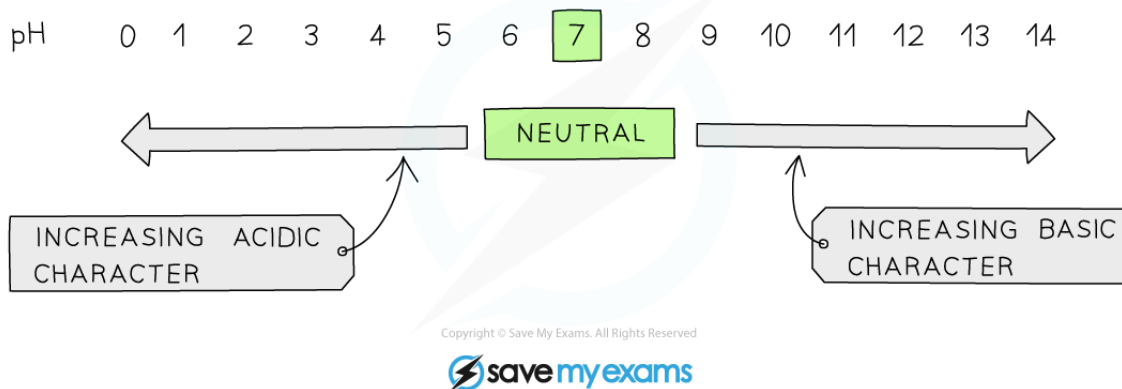


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8.2.3 Interpreting pH

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 H^+ is always **greater** than the concentration of 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 OH^- than H^+ ions
- Since the concentration of OH^- is always **greater** than the concentration of 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 OH^- and 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



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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 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:

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

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⁺] & [OH⁻] Table



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

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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**



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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 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 $1 \times 10^{-3} \text{ mol dm}^{-3}$:

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

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

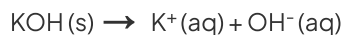
The final pH is 5, which corresponds to $1 \times 10^{-5} \text{ mol dm}^{-3}$

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

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

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

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

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



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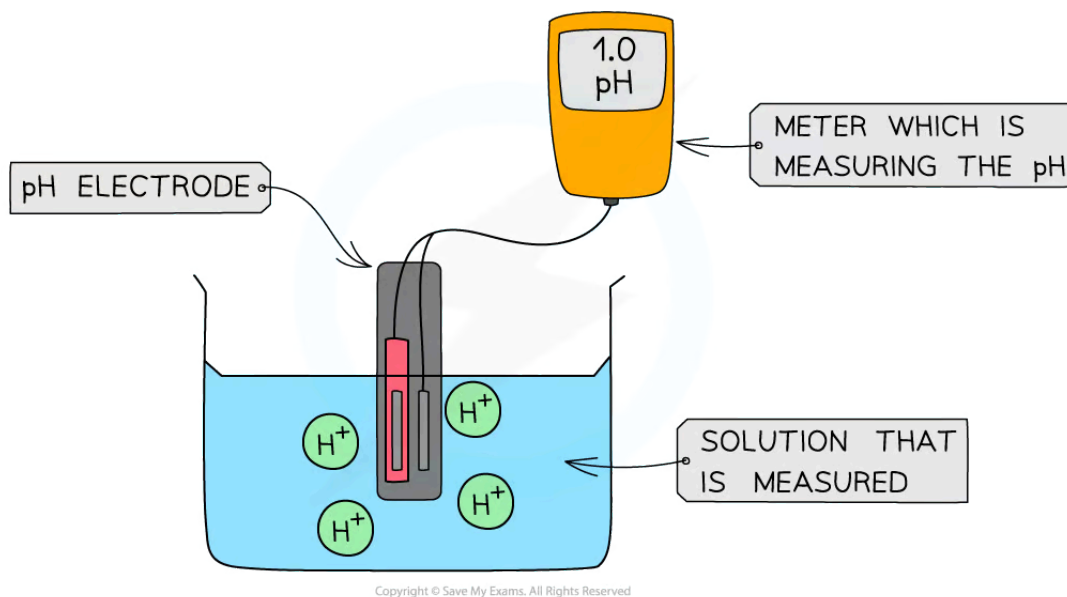


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8.2.6 pH Meters & Universal Indicator

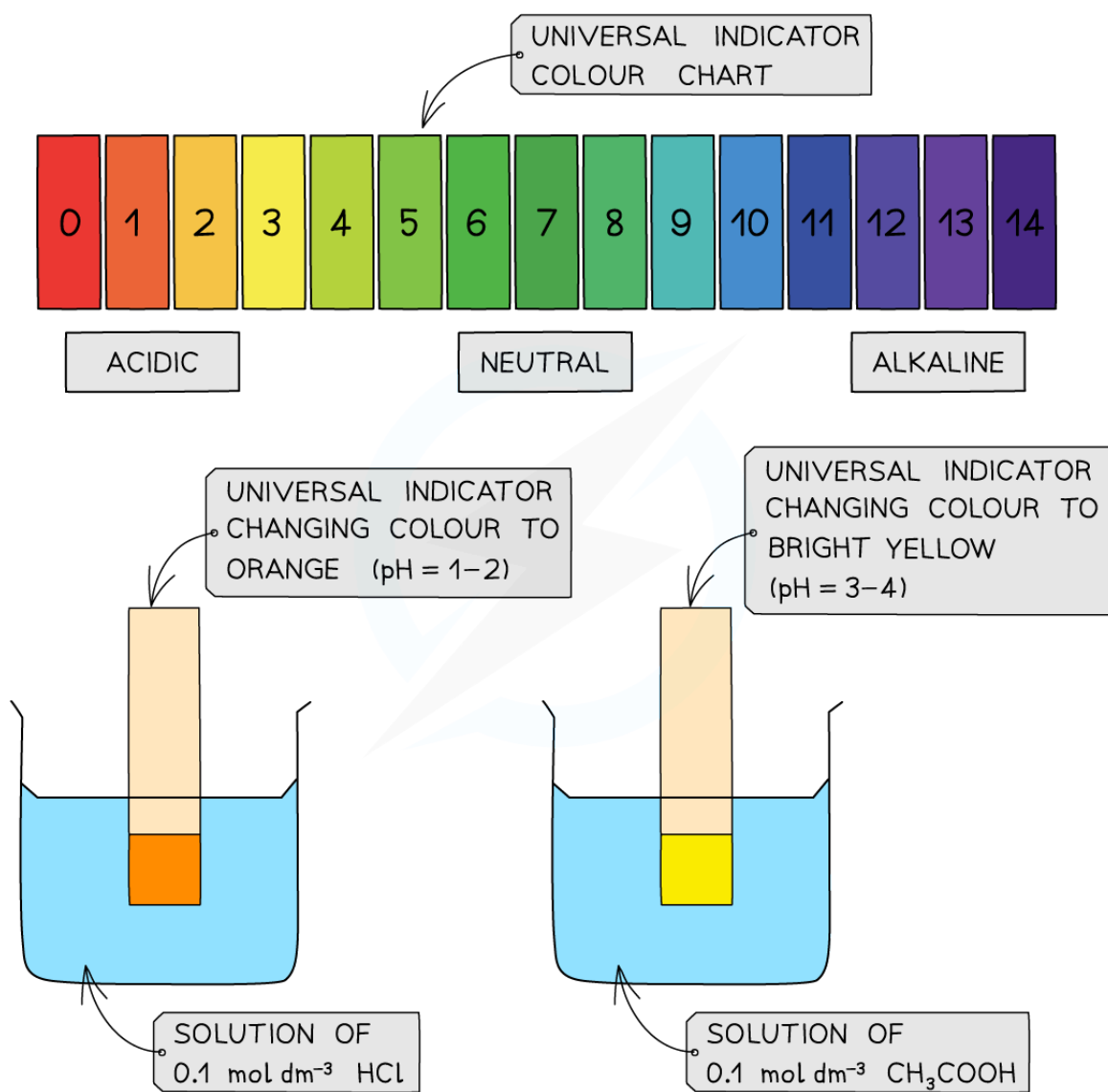
pH Meters & Universal Indicator

- 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 (CH₃COOH) acid. The colour chart is used to read off the corresponding pH values which are between 1–2 for HCl and 3–4 for CH₃COOH



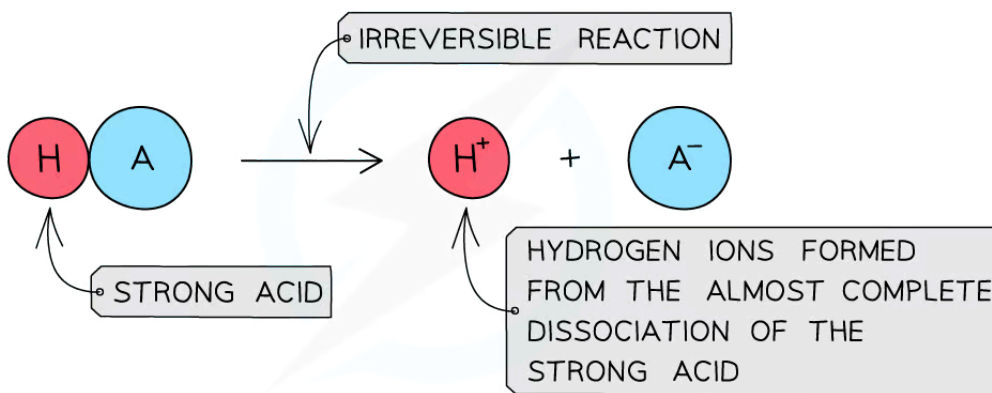
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8.2.7 Strong & Weak Acids & Bases

Strong & Weak Acids & Bases

Strong acids

- A **strong acid** is an acid that **dissociates** almost **completely** in aqueous solutions
 - HCl (hydrochloric acid), HNO₃ (nitric acid) and H₂SO₄ (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⁺/H₃O⁺ ions
- Since the **pH** depends on the concentration of H⁺/H₃O⁺ ions, the pH can be calculated if the concentration of the strong acid is known

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

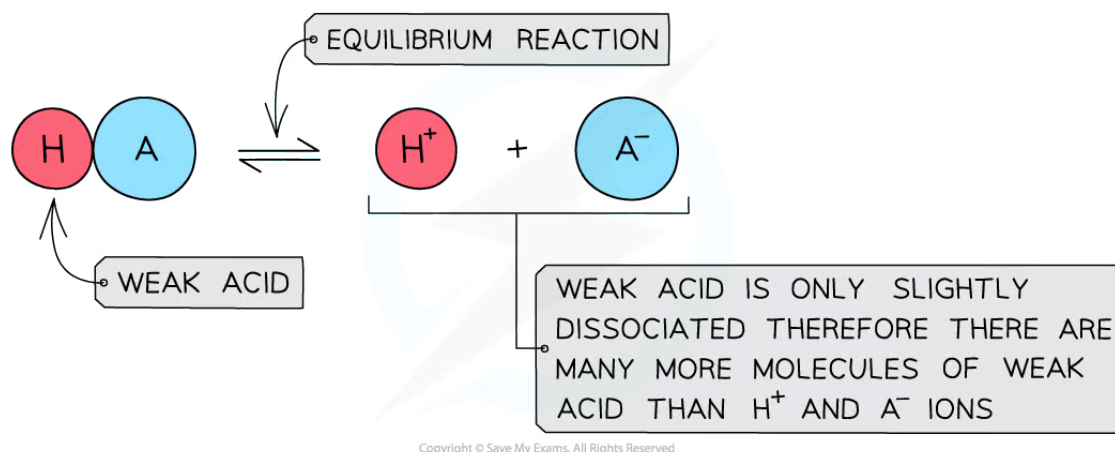
$$[\text{H}^+ (\text{aq})] = \text{CONCENTRATION OF H}^+/\text{H}_3\text{O}^+ \text{ IONS}$$

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pH is the negative log of the concentration of H⁺/H₃O⁺ 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), H₂S (hydrogen sulfide) and H₂CO₃ (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⁺/H₃O⁺ 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



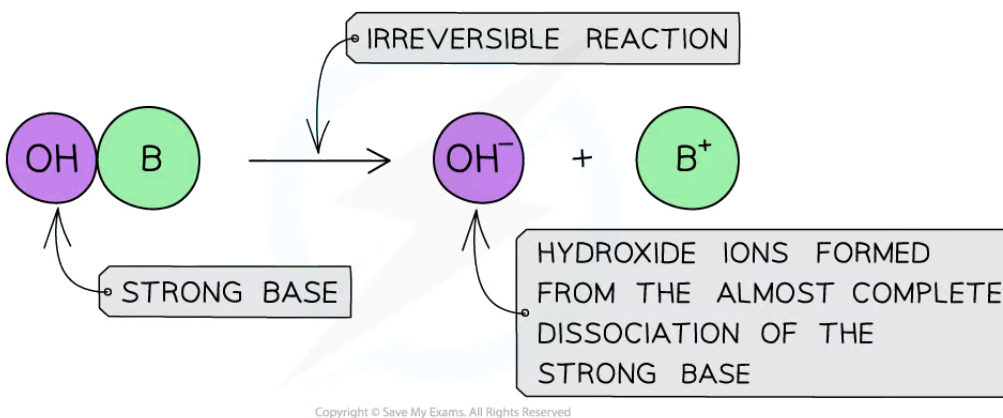
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	Strong Acid	Weak Acid
Position of Equilibrium	Right	Left
Dissociation	Completely (\rightarrow)	Partially (\rightleftharpoons)
H^+ concentration	High	Low
pH	Use [strong acid] for $[H^+]$	Use K_a to find $[H^+]$
Examples	HCl HNO ₃ H ₂ SO ₄ (first ionisation)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

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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

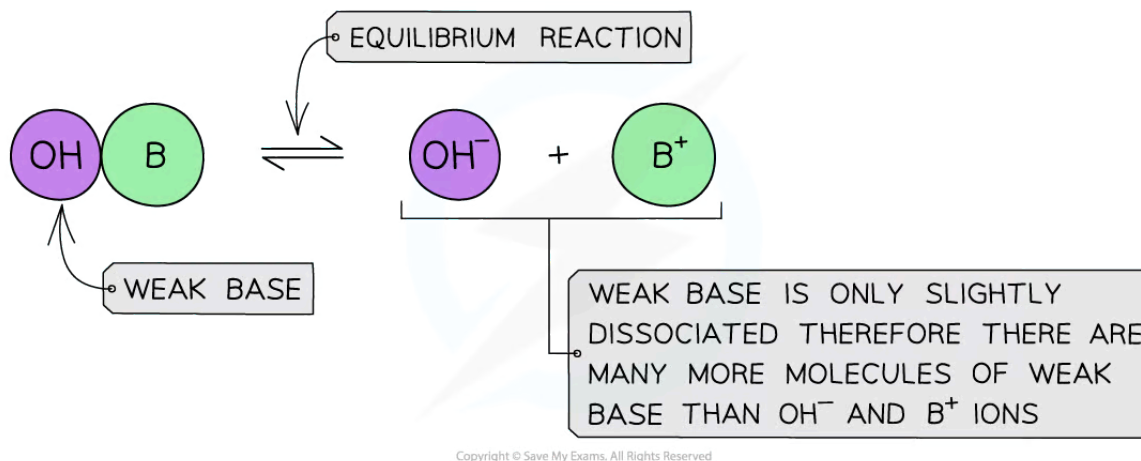


The diagram shows the complete dissociation of a strong base in aqueous solution

- The solution formed is highly basic due to the high concentration of the 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



Your notes

	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

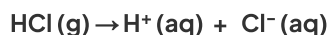
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Conjugate Pairs & Acid-Base Strength



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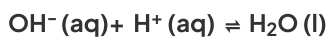
- The conjugate base of HCl is the chloride ion, Cl^- , but since the reverse reaction is virtually non-existent the chloride ion must be a very weak conjugate base



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_3O^+ or as H^+ however, if H_3O^+ is used, H_2O should be included in the chemical equation: $\text{HCl(g)} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ OR $\text{HCl(g)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ 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: $\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{SO}_4^{2-}$ HSO_4^- acts as a weak acid: $\text{HSO}_4^- \rightleftharpoons \text{H}^+ + \text{SO}_4^{2-}$ 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**.



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

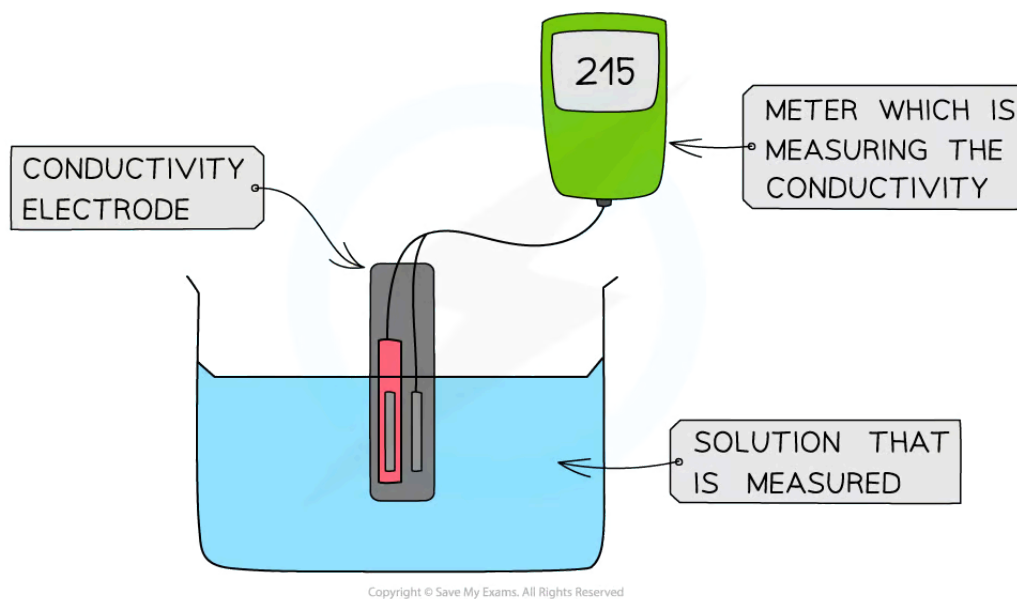
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Electrical conductivity

- 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



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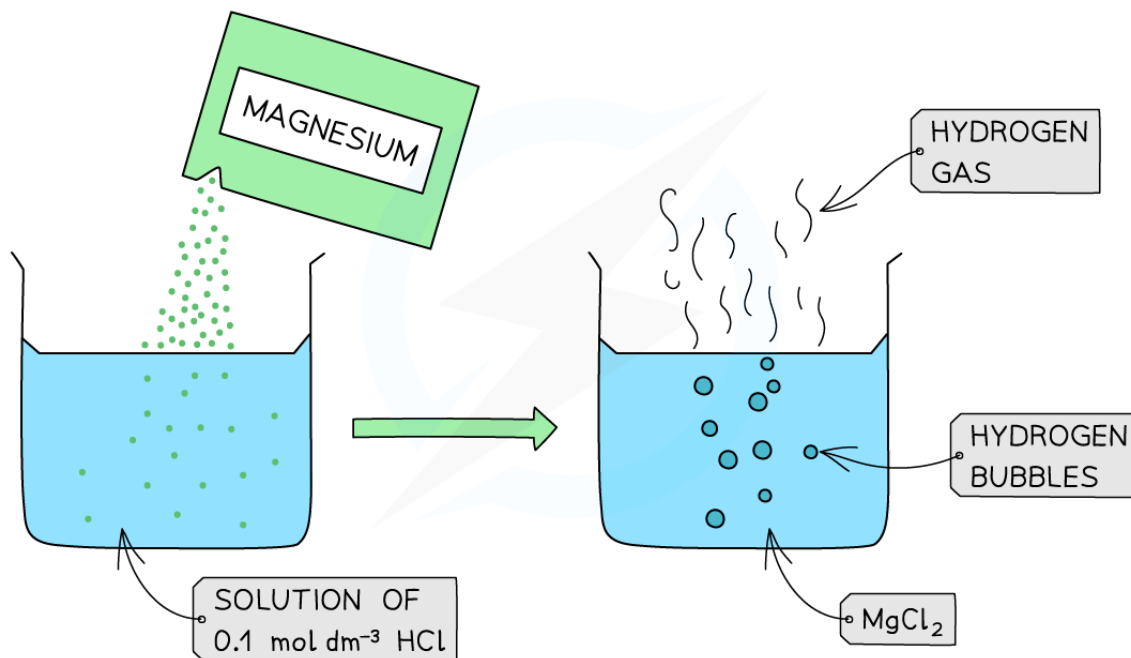
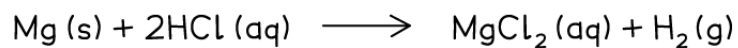
The diagram shows a digital conductivity meter that measures the electrical conductivity of a solution using an electrode

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

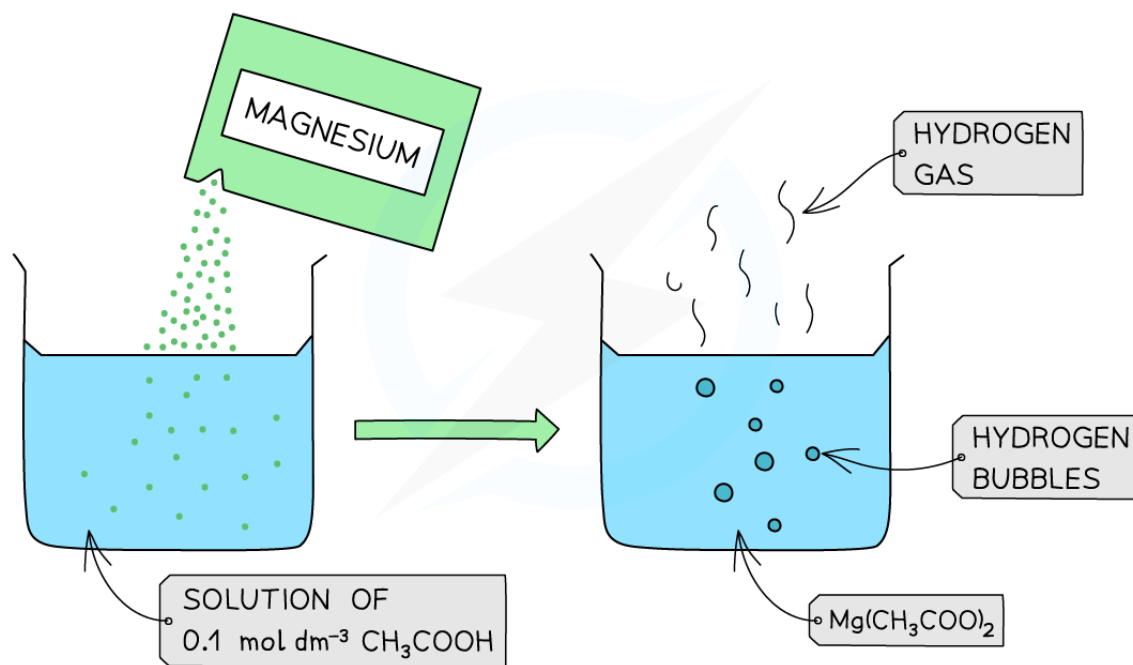
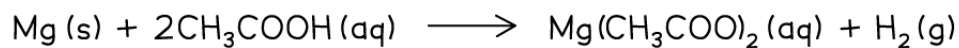


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The diagram shows the reaction of 0.1 mol dm^{-3} of a strong acid (HCl) with Mg. The reaction produces a lot of bubbles and hydrogen gas due to the high concentration of H^+ present in solution



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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.