

Edexcel Chemistry A-Level

Topic 12: Acid-base Equilibria

Detailed Notes





Brønsted-Lowry Acids and Bases

Acid-base equilibria involve the **transfer of protons** between substances. Therefore substances can be classified as acids or bases depending on their interaction with protons.

A Brønsted-Lowry **acid** is a **proton donor**. For example, Ammonium ions (NH_4^+).

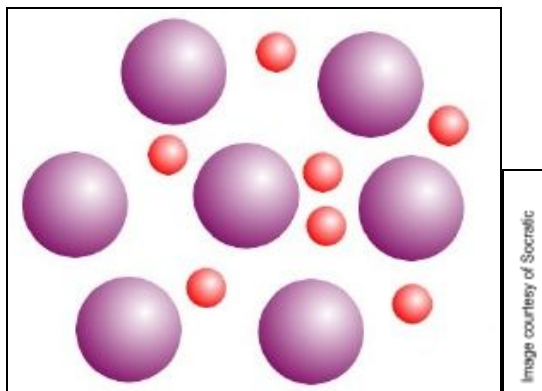
A Brønsted-Lowry **base** is a **proton acceptor**. For example, Hydroxide ions (OH^-).

Acid and Base Strength

Acid strength doesn't refer to the concentration of a solution. A **strong acid** is defined as being:

An acid that completely dissociates to ions when in solution with pH 0-1.

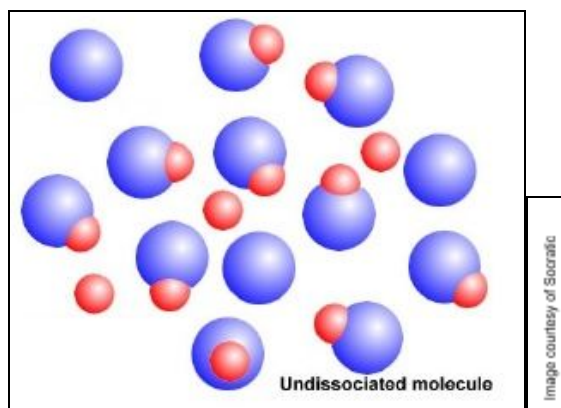
Example:



In comparison, a **weak acid** is defined as being:

An acid that only slightly dissociates when in solution with pH 3-5.

Example:



The same definitions are true for strong and weak bases. Strong bases have pH 12-14 and weak bases pH 9-11.





Determining pH

pH is a measure of **acidity and alkalinity**. It is a **logarithmic** scale from 0 to 14 giving the concentration of **H⁺ ions** in a solution. 0 is an **acidic** solution with a high concentration of H⁺ ions whereas 14 is a **basic** solution with a low concentration of H⁺ ions.

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

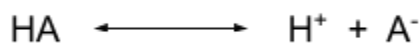
This equation also allows the concentration of H⁺ ions to be determined if the pH is known.

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

This concentration of H⁺ ions is equivalent to the **concentration of a strong acid** as it **completely dissociates** to ions in solution.

Weak Acids and Bases

Weak acids and bases only **slightly dissociate** in solution to form an **equilibrium** mixture. Therefore, the reaction has an equilibrium dissociation constant, **K_a**.



$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

In a similar way to [H⁺], this constant can be found using pK_a:

$$\text{pK}_a = -\log_{10} K_a$$

$$K_a = 10^{-\text{pK}_a}$$





These relationships of K_a , pK_a and $[H^+]$ can be used to find the **pH of weak acids and bases**. Depending on the reaction and the relative concentrations, a different method has to be used:

HA in excess - Use $[HA]$ and $[A^-]$ along with K_a to find $[H^+]$, then pH.

A^- in excess - Use K_w to find $[H^+]$, then pH.

HA = A^- - In this case, pK_a is equal to pH, therefore find pK_a .

Enthalpy Change of Neutralisation

When **strong** acids and bases react together they will produce **very similar enthalpies of neutralisation**. This is because the solutions completely dissociate so the same, simple acid-base reaction occurs between **H^+ and OH^- ions** to produce water in each case. The other dissociated ions present are simply **spectator ions** and do not affect the reaction.

However in reactions of **weak** acids and bases, the ions only slightly dissociate so **other enthalpy changes** also occur within the solution. As a result, the enthalpies of neutralisation can **vary** quite a lot.

Ionic Product of Water

Water **slightly dissociates** to ions as an equilibrium with its own equilibrium constant, **K_w** .

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

At **$25^\circ C$** , room temperature, K_w has a constant value of **1.0×10^{-14}** . However as temperature changes, **this value changes**.

The **forward** reaction in the equilibrium of water is **endothermic** and is therefore favoured when temperature of the water is increased. As a result, **more H^+ ions** are produced meaning the water becomes **more acidic as temperature increases**.



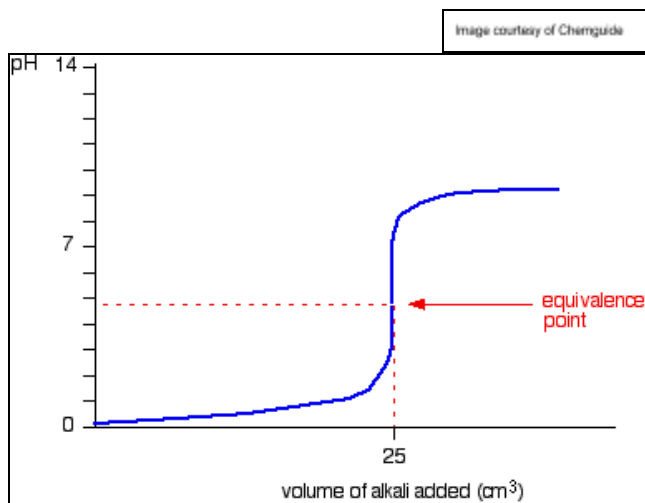


Titration Curves

A pH titration curve shows how pH of a solution **changes** during an **acid-base reaction**. When they react, a **neutralisation point** is reached, identified as a large **vertical section** of the graph, through the neutralisation or equivalence point.

To investigate, alkali is slowly added to an acid and the pH measured with a **pH probe** or vice versa. The **smaller** the added volumes, the **more accurate** the curve produced.

Example:



For a strong acid - strong base reaction, this neutralisation point occurs around **pH 7**. Other combinations of strong and weak acids and bases results in a **different neutralisation point**:

Strong Acid + Strong Base = pH 7

Strong Acid + Weak Base = < pH 7 (more acidic)

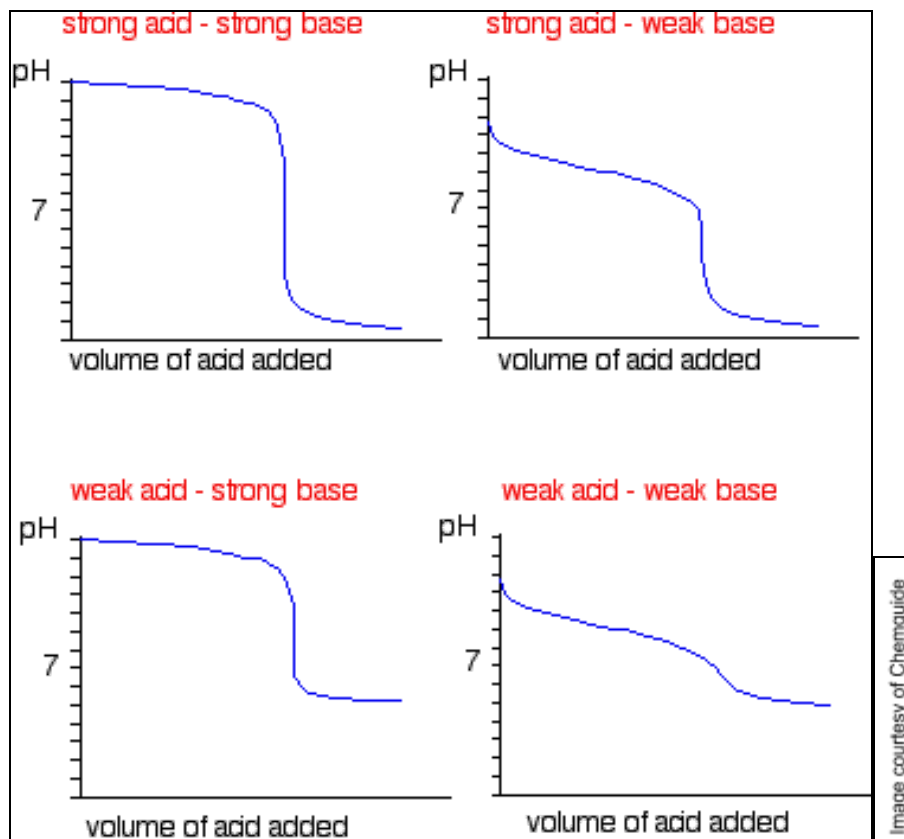
Weak Acid + Strong Base = > pH 7 (more basic)

Weak Acid + Weak Base = normally pH 7 but hard to determine





Example:



Indicators

Specific indicators have to be used for specific reactions as they can **only indicate a pH change within a certain range**.

The two most common indicators used at A-Level are **methyl orange** and **phenolphthalein**:

Methyl Orange - used for reactions with a **more acidic** neutralisation point.
- red in acids and turns **yellow** at the neutralisation point.

Phenolphthalein - used for reaction with a **more basic** neutralisation point.
- pink in alkalis and turns **colourless** at the neutralisation point.

Image courtesy of youbasic.weebly.com

The colours of indicators in acidic and basic solutions			
Indicator	Colour on acid side	pH at colour change	Colour on basic side
methyl orange	red	3–5	yellow
litmus	red	5–8	blue
phenolphthalein	colourless	8–10	pink

It is therefore important that the **correct indicator is selected** to use in a titration depending on the chemicals being used.

Buffer Action

A buffer solution is formed from an **acid and its salt** or an **acid and a base**. This produces a mixture containing **H⁺ ions** and a large **reservoir of OH⁻ ions** which helps to resist any change in pH. Therefore a buffer solution is defined as:

A solution which is able to resist changes in pH when small volumes of acid or base are added.

The large reservoir of OH⁻ ions allows the **ratio of acid to base** in the mixture to be kept almost **constant**.

Buffer Calculations

These are long calculations that use acid-base calculations. There are two types:

- Acid + Base** - Find the number of moles of each species.
- Calculate their concentration when at equilibrium using the total volume.
 - Use K_a to find $[H^+]$ and therefore pH.

- Acid + Salt** - Find the moles of the salt.
- Use K_a to find pH.

Adding small volumes

The pH of a buffer solution doesn't change much but will change in the **order of 0.1 or 0.01 units of pH** when a **small volume** of acid or base is added.

Adding small amounts of **acid (H⁺) increases the concentration** of the acid in the buffer solution meaning the overall solution will get slightly **more acidic**.

Adding small amounts of **base (OH⁻) decreases the concentration** of acid in the buffer solution meaning the overall solution will get slightly more **basic**.

Uses of Buffers

Buffer solutions are common in **nature** in order to keep systems regulated. This is important as **enzymes** or reactions in living organisms often require a **specific pH**, which can be maintained using a buffer solution.





Another important buffer in nature is found in the human circulatory system. The **pH of human blood** is maintained in a buffer between **carbonic acid and bicarbonate ions**. These ions **neutralise any acidic substances** that enter the bloodstream, converting them to carbonic acid and water.

Example:

