AQA Chemistry A-level

3.1.12: Acids and Bases

Detailed Notes

3.1.12.1 - 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**.

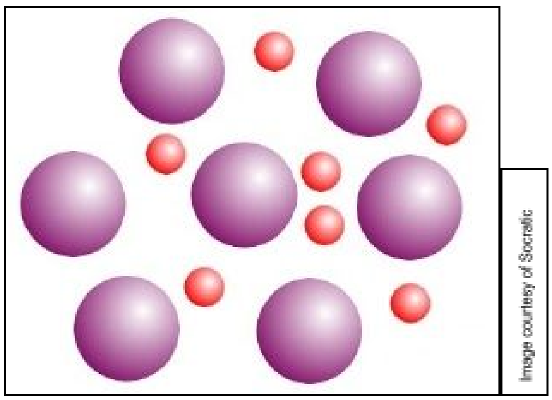
A Brønsted-Lowry **acid** is a proton **donor**. For example, Ammonium ions (NH_4^+).
A Brønsted-Lowry **base** is a proton **acceptor**. For example, Hydroxide ions (OH^-).

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 3-5.

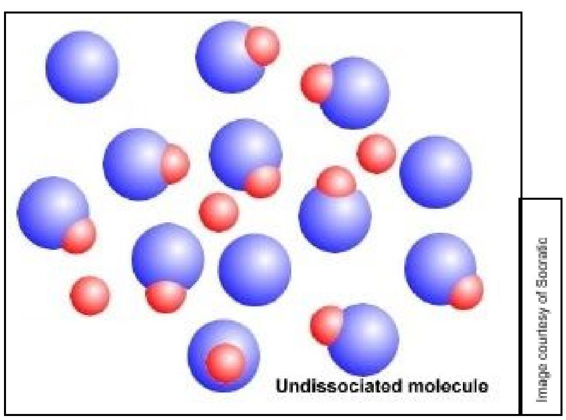
Example:



In comparison, a weak acid is defined as being:

An acid that only slightly dissociates when in solution with pH 0-1.

Example:



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

3.1.12.2 - Determining pH

pH is a **measure of acidity and alkalinity**. It is a l**ogarithmic scale** from 0 to 14 giving the concentrations 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.

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

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

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

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

3.1.12.3 - Ionic Product of Water

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

$$\boxed{Kw = [H^+][OH^-]}$$

At 25°C, room temperature, Kw has a constant value of 1×10^{-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**.

$$\boxed{H_2O \longleftrightarrow H^+ + OH^-}$$

3.1.12.4 - 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, Ka.

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

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

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

$$\boxed{pKa = -\log_{10}Ka}$$

$$\boxed{Ka = 10^{-pKa}}$$

These relationships of Ka, pKa 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 Ka to find $[H^+]$, then pH.

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

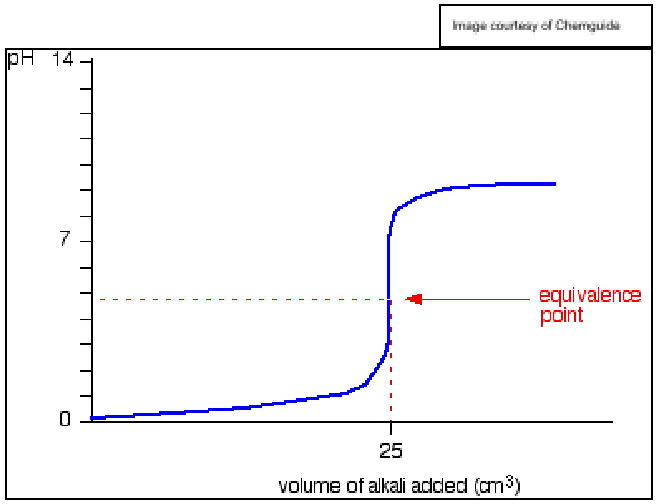
$HA = A^-$ - In this case, pKa is equal to pH, therefore find pKa.

3.1.12.5 - 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 which is identified as a **large vertical section** 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:



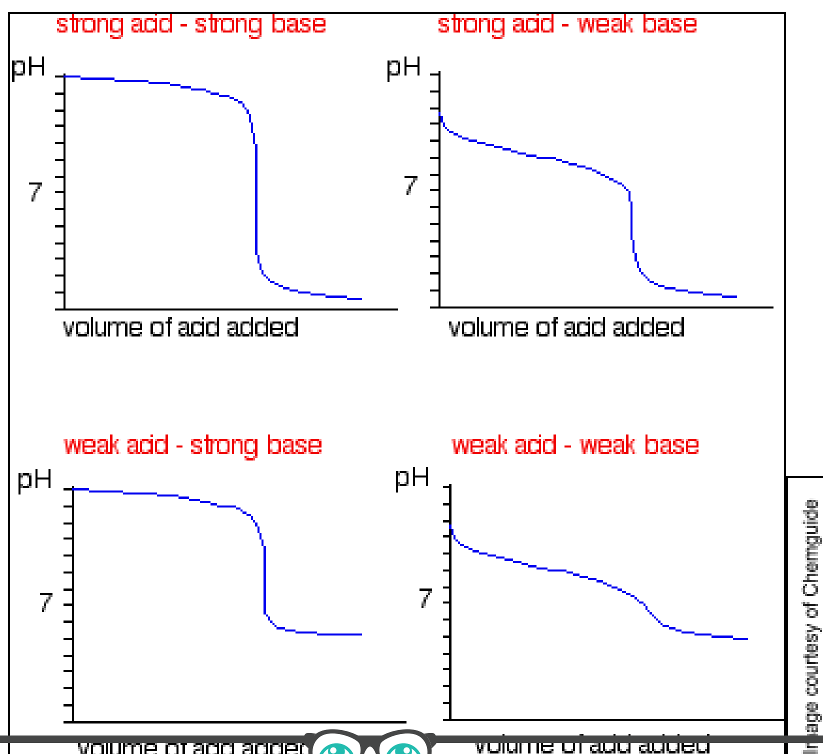
For a strong acid - strong base reaction, this neutralisation point occurs around pH 7. Other combinations of strong and weak acids and bases results in a different neutralisation point:

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



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.
- orange 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.

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

Image courtesy of youbasic.weebly.com

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

3.1.12.6 - Buffer Action

Acidic buffer solutions contain a **weak acid** and the **salt of that weak acid** and basic buffer solutions contain a **weak base** and the **salt of that weak base**.

If we take an example of Ethanoic acid and Sodium Ethanoate, if an acid (H^+) is added it will resist a change in pH by reacting the ethanoate ions with the H^+ to make Ethanoic Acid. As the H+ ions are removed there is little to no 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.

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 Ka to find [H+] and therefore pH.

Acid + Salt - Find the moles of the salt.
- Use Ka 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.