WJEC Chemistry A-Level

C2.1: Simple Equilibria and Acid-Base Reactions

Detailed Notes
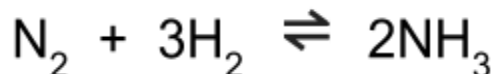
English Specification

Reversible Reactions

Some reactions are **reversible**, where the reactants react to form products, which can then react to reform the original reactants. It is represented using a **double arrow**.
Example:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

In each reaction system, one reaction will be **endothermic** and the opposite reaction will be **exothermic**.

Dynamic Equilibrium

A reversible reaction reaches dynamic equilibrium when the **rate of the forward reaction** equals the **rate of the backwards reaction**. At equilibrium, the **concentration** of reactants and products remains constant.

Dynamic equilibrium can only occur in **closed systems** where no substances can get in or out in order to influence the reactions.

Le Chatelier's Principle

This theory is used to predict how an equilibrium mixture will change if the reaction conditions are altered. It states:

**When a system is subject to change, the system will
alter to lessen the effect of that change.**

It is a useful principle to consider in industry, so that reaction conditions are chosen to **maximise the yield** of a desired product.

Changing Temperature

Increasing temperature favours the **endothermic reaction** (+ve ΔH) as the **excess heat** needs to be removed from the system to lessen the effect of the initial increase. Therefore it will **increase the yield of the products in the endothermic direction**.
Likewise, decreasing temperature favours the exothermic reaction (-ve ΔH) as heat needs to be gained and so the **yield of the products in the exothermic direction** will be increased.

In exam questions, the ΔH of the forward reaction will be given so it is clear which direction is endothermic and which is exothermic.

Changing Pressure

Increasing pressure favours the side of the reaction with **fewer moles** as this will help to **release the build up in pressure**. It will increase the yield of the products on this side of the reaction.

Likewise, decreasing pressure favours the side of the reaction with **more moles** as pressure has been lost. The yield of the products on this side of the reaction will be increased.

Changing Concentration

Increasing the concentration of the reactants will favour the forward reaction that produces the products as more molecules of reactant are available to react. Therefore it will **increase the yield of the products**.

Likewise, increasing the concentration of products will favour the backwards reaction that **reforms the reactants**.

Catalysts

Catalysts appear not to affect the equilibrium position as they **increase the rate of the forward and backward reaction equally**. However, catalysts do allow equilibrium to be reached **faster**.

The Equilibrium Constant (Kc)

For reversible reactions, there is an **equilibrium constant**, **Kc**, that indicates the **position of equilibrium** for a reaction at a certain temperature. Kc for a reaction always has the **same value** unless the reaction conditions, and therefore the position of equilibrium, are changed.

In order to find the value of Kc, the **concentrations** of the substances in the reaction at the **equilibrium position** need to be obtained. Kc can be **calculated** as follows:
Example:

$$aA + bB \rightleftharpoons cC + dD$$

$$Kc = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$
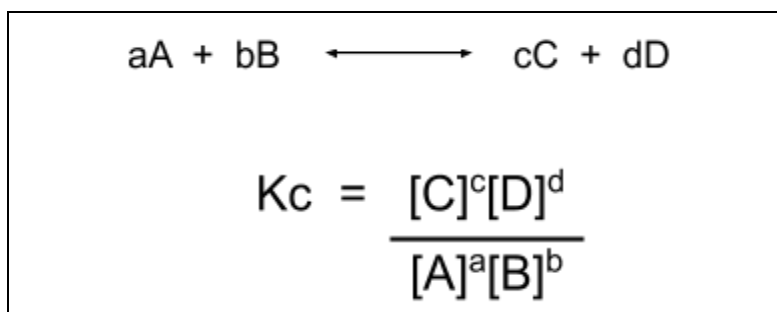
Using this principle, equilibrium calculations can be carried out to determine the quantities of the species present in a reaction.

The equilibrium constant has **varying units** depending on the chemical reaction. It can be calculated by **substituting the units** into the Kc expression. Some of these then **cancel** to give the overall units of Kc for that reaction.

Example:

$$\frac{moldm^{-3}}{(moldm^{-3})^2\ moldm^{-3}} = \frac{1}{mol^2dm^{-6}}$$

$$= mol^{-2}dm^6$$

The value of Kc is **not affected** by **concentration change, pressure change** or use of a **catalyst**. Kc is only affected by changing the reaction **temperature** as the equilibrium position will change, resulting in different concentrations of reactants and products.

Acids and Bases

Acid-base equilibria involve the **transfer of protons** (H^+ ions) between substances. Therefore, substances can be classified as **acids or bases** depending on their interaction with the hydrogen ions.

An **acid** is a **donor** of H^+ ions. For example, ammonium ions (NH_4^+).
A **base** is an **acceptor** of H^+ ions. For example, hydroxide ions (OH^-).

Determining pH

pH is a measure of **acidity and alkalinity**, related to the concentration of H^+ ions present. It is a **logarithmic 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.

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

These equations show the relationship between pH and H^+ ions.

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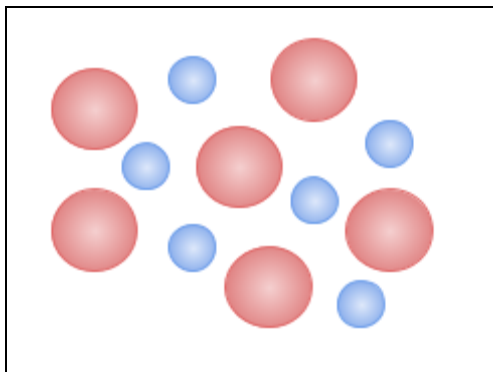
Acid and Base Strength

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

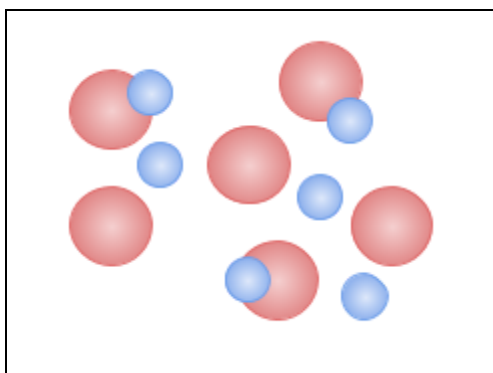**An acid that (almost) completely dissociates into its ions.**

Example:

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

An acid that only slightly dissociates into its ions.

Example:

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

Acid-Base Titrations

A **titration** is a practical method used to determine the **unknown concentration** of an acid or base by finding its **neutralisation** point. A simple titration can be carried out between a strong acid and a strong base.

Indicators are used to help determine the neutralisation point of the reaction. 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.

Indicator	Colour in Acid	pH at Colour Change	Colour in Base
Methyl Orange	red	3-5	yellow
Litmus	red	5-8	blue
Phenolphthalein	colourless	8-10	pink

pH can be measured and recorded throughout the reaction and used to plot a **pH titration curve.** These curves show how pH of a solution changes during an acid-base reaction.

When the species react, a neutralisation point is reached which is identified as a **large vertical section** on the pH curve.
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:

pH

Equivalence Point

Volume of Alkali Added / cm^{-3}
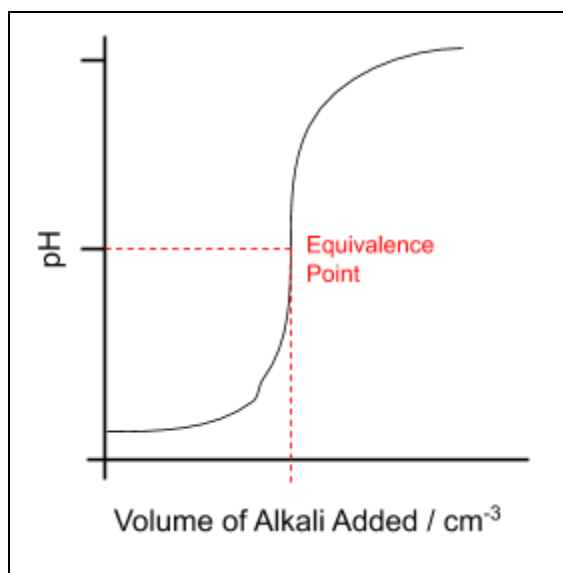
For a strong acid - strong base reaction, this neutralisation point occurs around **pH 7**. Other combinations of strong and weak acids and bases result in different neutralisation points:

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