# 7.1 – Dynamic Equilibrium

7.1.1 - Outline the characteristics of chemical and physical systems in a state of equilibrium

Open system – When a reaction occurs in an unsealed container

Closed system – When a reaction occurs in a sealed container, and it is possible to reach equilibrium.

When a reaction takes place, the products are then able to revert back to their reactants, so that the reaction is occurring **both forwards and backwards**. At the beginning of any reaction, the particles in the highest concentration are the reactants. As more of the products are formed, this concentration decreases, while the concentration of the product increases.
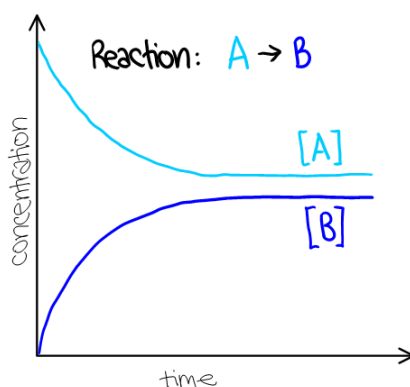
Since the concentration of reactants decreases, this also decreases the probability of a collision between the particles, thus slowing the **rate of reaction**. The inverse is also true for the products.
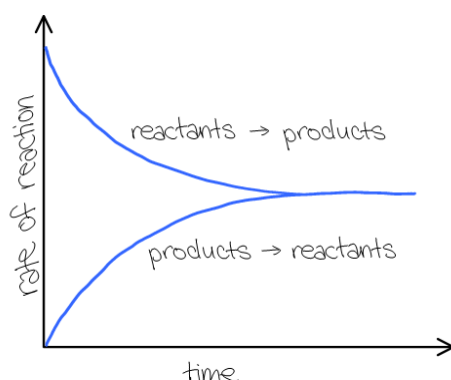
Looking back at the first graph, we can see that the reactants and products reach a point in time when the reaction appears to have stopped. However, as we see in the second graph, the reactions do not cease, but in fact are continuing at the same rate. All the macroscopic properties of the reaction will be constant and has reached a state of **dynamic equilibrium**.

The characteristics of dynamic equilibrium are:

- The forward and reverse reactions are happening at the **same rate**
- Occurs for physical systems also (i.e. water vapour changing state to liquid water)
- Depending on factors such as temperature and pressure, the reaction may favour the products or the reactants.

7.2 – The Position of Equilibrium

7.2.1 - Deduce the equilibrium constant expression (K_c) from the equation for a homogeneous reaction

A **homogenous reaction** is when the reactants and the products are all in the same state. The equilibrium law shows that the concentration of reactants and products have a relationship. It is expressed:

$$aA + bB + \cdots \leftrightarrow pP + qQ + \cdots$$

The coefficients in the equation above are important. The relationship between them to calculate the equilibrium constant is shown below:

$$K_c = \frac{[P]^p \, [Q]^q \, \ldots}{[A]^a \, [B]^b \, \ldots}$$

Note that this is **dependent of temperature**. The square brackets mean the concentration of each species.

Equilibrium constants will have different units. It is also possible to calculate this for the pressure of a gaseous reaction, labelled K_p.

7.2.2 - Deduce the extent of a reaction from the magnitude of the equilibrium constant

$$K_c \gg 1 - the\ reaction\ goes\ almost\ to\ completion$$
$$K_c \ll 1 - the\ reaction\ hardly\ proceeds$$

When the **constant becomes larger**, this means that the reaction favours the right hand side of the equation and more of the products are made. When the **constant becomes smaller**, the reaction favours the left-hand side of the equation and fewer of the products are made.

7.2.3 - Apply Le Chatelier's principle to predict the qualitative effects of changes in temperature, pressure and concentration on the position of equilibrium and on the value of the equilibrium constant

The equilibrium constant changes according to **temperature**. Le Chatelier showed that when conditions change in the environment of a reaction, then the reaction will move the equilibrium constant to accommodate it.

Temperature Change	Exothermic Reaction		Endothermic Reaction	
	Position of Equilibrium	Equilibrium Constant, K_c	Position of Equilibrium	Equilibrium Constant, K_c
Increase	Moves to the left	Decrease	Moves to the right	Increase
Decrease	Moves to the right	Increase	Moves to the left	Decrease

The <u>position of equilibrium</u> is also affected by **concentration**, however the <u>equilibrium constant</u> is not.

Change in Concentration	Position of equilibrium of reactants	Position of equilibrium of products
Increase	Moves to the right	Moves to the left
Decrease	Moves to the left	Moves to the right

Le Chatelier's principle also applies to **pressure** of gases, in that the <u>position of equilibrium</u> will change, but the <u>equilibrium constant</u> does not.

Change in Pressure	Reaction where n(reactants) > n(products)	Reaction where n(reactants) < n(products)
Increase	Moves forwards to the right	Moves backwards to the left
Decrease	Moves backwards to the left	Moves forwards to the right

Note that for the **dilution** of a reaction taking place in solution, the reaction will favour the side with the highest number of moles of particles in order to increase the concentration of all the particles.

7.2.4 - State and explain the effect of a catalyst on an equilibrium reaction

Catalysts increase the rate of reaction so that **equilibrium is reached faster** with changing the equilibrium constant or the position of equilibrium. Eventually, the reactants and products will still reach the same concentrations as if a catalyst were not used.
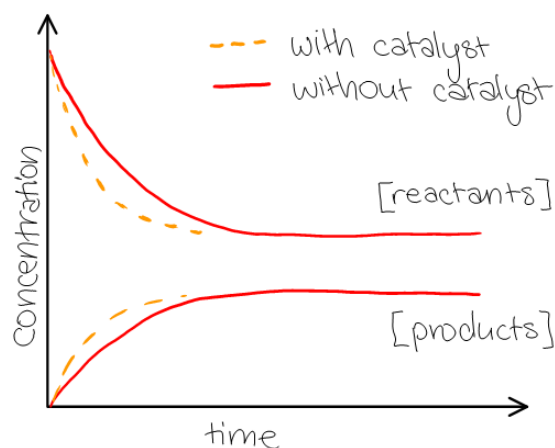
7.2.5 - Apply the concepts of kinetics and equilibrium to industrial processes

When considering the factors for the industrial production of a chemical substance, the following must be taken into account:

- **Rate of reaction** – How appropriate is it and is it high enough to make it useful?
- **Yield** – Will a sufficient amount of the substance be produced?
- **Economics** – How much will it cost to create the optimum conditions and will the revenue outweigh this?

The producers will try to optimise all these things so that they can produce the substance quickly, in large quantities and for the lowest cost possible.

The Haber Process

The Haber process is used for the industrial production of **ammonia** which is used for:

- Fertiliser
- Explosives
- Nitric Acid
- Polymers

$$N_{2\,(g)} + 3H_{2\,(g)} \leftrightarrow 2NH_{3\,(g)} \qquad \Delta H = -92.4\ kJ\ mol^{-1}$$

The principles of equilibrium can be used to increase the rate of reaction and yield, such as increasing the concentration of the reactants, removing products as they form, increasing gas pressure or decreasing the temperature (as the forward reaction is exothermic).

Real conditions used:

- **Pressure**: 100-350 atm
  - Lower pressure used to ensure safety
  - More expensive at such a high pressure, but increases the rate and yield
- **Temperature**: 400-450°C
  - Although the yield would be greater at lower temperatures, the temperature is increased to speed up the rate of reaction.
- **Catalyst**: Porous iron (Fe)
  - Cheap and effective
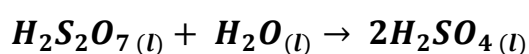 - Reactants are passed over the iron several times until a significant yield is produced
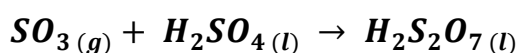
The Contact Process

The Contact process produces sulfuric acid. However, the reaction takes place in four steps, the second of which is an equilibrium reaction.

$$2SO_{2\,(g)} + O_{2\,(g)} \leftrightarrow 2SO_{3\,(g)} \qquad \Delta H = -197\ kJ\ mol^{-1}$$

The following reactions are:

$$SO_{3\,(g)} + H_2SO_{4\,(l)} \rightarrow H_2S_2O_{7\,(l)}$$

$$H_2S_2O_{7\,(l)} + H_2O_{(l)} \rightarrow 2H_2SO_{4\,(l)}$$

Since one of the reactions is an equilibrium reaction, the rate of reaction and the yield can be increased by:

- Increasing the concentration of the reactants
- Removing product as it forms
- Increasing gas pressure
- Decreasing the temperature

With all this in mind, the actual conditions for the Contact process are:

- Excess air added
- Atmospheric pressure
- 440°C for incoming gases

A **catalyst** is also used to increase the rate of reaction – vanadium oxide, V_2O_5, which is laid out in pellets to increase surface area.