Unit 3 - Chemistry in Society

3D Equilibria

3D1 Dynamic equilibrium

Chemical reactions which take place in both directions are called reversible reactions. The following is an example of a reversible reaction - hydrogen and iodine reacting to form hydrogen iodide.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

A reversible reaction attains a state of equilibrium when the rate of the forward reaction is equal to the rate of the reverse reaction. In the equilibrium mixture both the forward and reverse reactions are taking place. Since the rates of the forward and reverse reactions are equal, the concentration of reactants and products remain constant, though rarely equal. The system is said to be at dynamic equilibrium.

Equilibrium can only be attained in a closed system - reactants or products cannot be removed.

Shifting the equilibrium position

The proportion of products to reactants in an equilibrium mixture is described as the equilibrium position.

$$A + B \rightleftharpoons C + D$$

If the conversion of A and B into C and D is small the position of equilibrium lies to the left, or to the side of the reactants. If the equilibrium mixture is largely composed of C and D, the position of equilibrium lies to the right, or to the side of the products.

3D2 Le Chatelier's Principle

For a system in equilibrium, alteration of one of the factors (pressure, temperature or concentration) will cause the position of equilibrium to shift to reduce the effects of the imposed conditions.

1. **Concentration**

$$Cr_2O_7^{2-}(aq) + OH^- \rightleftharpoons 2CrO_4^{2-}(aq) + H^+(aq)$$

dichromate (orange) chromate (yellow)

In the above reaction adding sodium hydroxide increases the concentration of OH^- ions and pushes the equilibrium to the right giving a yellow colour. Adding hydrochloric acid increases the concentration of H^+ ions and pushes the equilibrium to the left giving an orange colour.

2. **Pressure**

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

Gas pressure arises from the bombardment of gas molecules. The greater the number of molecules the greater the pressure.

Formation of nitrogen dioxide results in twice as many molecules. Increase in pressure will favour the equilibrium moving to the left (reducing the effects of pressure), and decrease in pressure will favour the equilibrium moving to the right.

3. **Temperature**

$$Co(H_2O)_6^{2+}(aq) + 4Cl^-(aq) \rightleftharpoons CoCl_4^{2-}(aq) + 4H_2O(l) \qquad \Delta H > 0$$

pink blue

The above reaction is endothermic so an increase in temperature will favour movement of equilibrium to the right. At higher temperatures the solution is blue.

The reverse reaction will be exothermic and the opposite is the case. The reaction is trying to get rid of heat so applying heat will hinder it. At lower temperatures the solution is pink.

4. **Catalyst**

A catalyst speeds up the rate of attainment of equilibrium but does not affect the position of equilibrium.

3D3 The Haber Process

Because many reactions are at equilibrium and do not go to completion the result may be that costly reactants fail to be completely converted into products.
To maximise profits, chemists employ strategies to move the position of equilibrium in favour of products.
The Haber process is the industrial process producing ammonia from nitrogen and hydrogen.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -91 \text{ kJ}$$

1. **Pressure**

An increase in pressure favours pushing the equilibrium to the right as there are less molecules.

2. **Temperature**

Being an exothermic reaction means that an increase in temperature will drive the reaction to the left, ie. a decrease in temperature will favour formation of ammonia. In practice low temperatures slow up the reaction, so higher temperatures are used.

3. **Catalyst**

A catalyst allows equilibrium to be reached more quickly. This means that lower temperatures can be used which will favour formation of ammonia.

A suitable balance of all three conditions is used

temperature	about 500°C
pressure	about 150 to 300 atm
catalyst	iron

4. **Recycling of unreacted gases and the removal of product**

Removal of ammonia by condensing makes the equilibrium continually alter to the left, so equilibrium is never attained. Unreacted nitrogen and hydrogen are recycled so that although conversion to ammonia is low the process is efficient.