<u>CHEMICAL REACTIONS AND ENERGY CHANGES (ENERGETICS)</u>

Energy is the ability to do work. Chemicals possess energy in two ways:

1. Kinetic energy
2. Chemical energy

During chemical reactions, bonds are broken in the reactants. Energy is absorbed and this is an <u>endothermic</u> process. Reactants need a minimum amount of energy for the reaction to proceed; this is known as the <u>activation energy.</u>

Additionally, in chemical reactions, new bonds are formed in order to yield products. Energy is released in the formation of new bonds and this is known as an <u>exothermic</u> process. These energy changes occur together, and are measured as changes in heat energy.

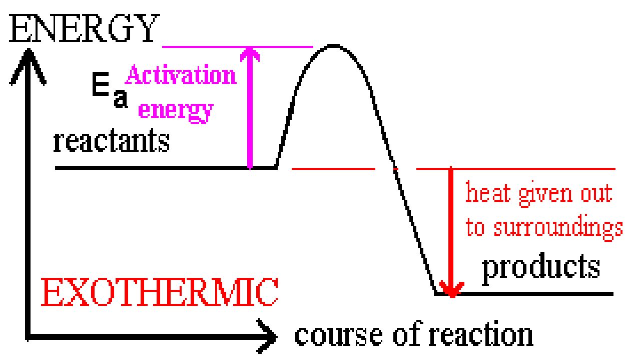
Overall, a reaction is:

- **Endothermic** if the energy that is absorbed in breaking bonds is greater than the energy released in the forming of new bonds. This means that heat energy is absorbed from the surroundings, and the reaction becomes cooler. E.g. i) Dissolving ammonium chloride and potassium nitrate in water; ii) Photosynthesis
  Endothermic reactions are much rarer than exothermic reactions.

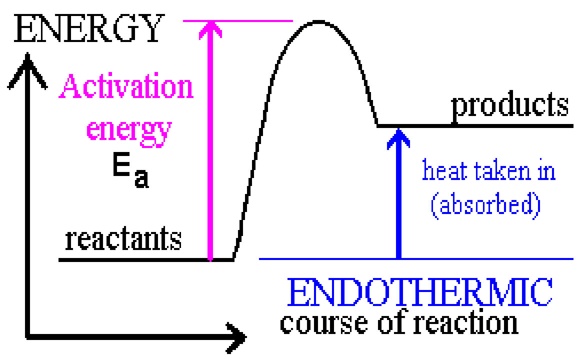
- **Exothermic** if the energy absorbed in breaking bonds is less than the energy released in forming new bonds. As a result, heat energy is released to the surroundings, and the reaction becomes hotter. E.g. i) Respiration; ii) Neutralization reactions; iii) Combustion reactions; iv) Dissolving ethanol in water

The energy or heat content of any substance is called its enthalpy (H); change in enthalpy is written as ΔH.

$$\begin{aligned}&Enthalpy\ Change\ of\ reaction\\&\qquad = Energy\ content\ of\ products - Energy\ content\ of\ reactants\end{aligned}$$



Energy Profile Diagram of an Exothermic Reaction



Energy Profile Diagram of an Endothermic Reaction

Measuring Heat Changes in the Lab

When measuring heat changes, the following apparatus are needed:

1. An insulated container which serves as a **calorimeter.** E.g. Styrofoam cups, plastic cups
2. Thermometer
3. A balance
4. A volumetric apparatus, E.g. beaker, measuring cylinder, burette, pipette

General Steps in the Procedure

1. Allow reactants to reach room temperature. Record the temperature.
2. Mix the reactants and record the highest or lowest temperature reached.
3. Determine the temperature change of the reaction.
4. Calculate the heat change for the reaction.

$$Heat\ Absorbed\ or\ Given\ Out = mass \times specific\ heat\ capacity \times temperature\ change$$

$$\Delta H = mc\Delta T\ or\ mc\Delta\theta$$

For aqueous solutions, 1 cm^3 is assumed to have a mass of 1 g. The specific heat capacity of these solutions would therefore be the sum of that of water which is **4.18 $Jg^{-1}{}^{\circ}C^{-1}$**.

Specific Heat Capacity

This is the amount of energy require to raise the temperature of 1 g of water by 1°C.

Heat of Neutralization

This is the energy change which occurs when 1 mole of water is formed during a neutralization reaction.

Heat of Combustion

This is the heat change which occurs when 1 mole of a substance is completely burnt in oxygen.

Heat of Solution

This is the heat change which occurs when 1 mole of a solute dissolves in a specific volume of solvent.

CALCULATIONS

Heat of Neutralization

100 cm^3 of 1 mol dm^{-3} NaOH solution at 18°C are added to 100 cm^3 of 1 mol dm^{-3} HCl solution at 18°C. The temperature of the resulting salt solution is 24.8°C. Calculate the heat of neutralization.

$$Increase\ in\ temperature = 24.8℃ - 18℃ = 6.8℃$$
$$Heat\ evolved = mc\Delta\theta = 200\ g \times 4.2\ Jg^{-1}℃^{-1} \times 6.8℃ = 5712\ J = 5.7\ kJ$$
$$Since\ the\ reaction\ is\ a\ neutralization\ reaction, it\ is\ exothermic, and\ therefore, the\ heat\ evolved = -5.7kJ$$

$$1000\ cm^3\ NaOH\ contains\ 1\ mole$$
$$\therefore\ 100\ cm^3 contains\ \frac{100}{1000} \times 1 = 0.1\ mole$$
$$1000\ cm^3\ HCl\ contains\ 1\ mole$$
$$\therefore\ 100\ cm^3 contains \frac{100}{1000} \times 1 = 0.1\ mole$$

$$NaOH + HCl \rightarrow NaCl + H_2O$$
$$1\ mole\ NaOH + 1\ mole\ HCl \rightarrow 1\ mole\ H_2O$$
$$\therefore\ 0.1\ mole\ NaOH + 0.1\ mole\ HCl \rightarrow 0.1\ mole\ H_2O$$

$$When\ 0.1\ mole\ H_2O\ is\ formed, -5.7\ kJ\ of\ heat\ is\ evolved$$
$$\therefore 1\ mole\ of\ H_2O\ will\ give\ \frac{1}{0.1} \times (-5.7) = -57kJ$$

The heat of neutralization is –57 kJ.

Heat of Solution

When 2.12 g anhydrous Na_2CO_3 is dissolved in 100 g of water, a rise in temperature of 1.17 °C is observed. Calculate the molar heat of solution of Na_2CO_3. (RAM: Na = 23, C = 12, O = 16)

$$Heat\ evolved = mc\Delta\theta = 100\ \mathrm{g} \times 4.2\ J\ g^{-1}℃^{-1} \times 1.17℃ = 491.4\ J$$
$$2.12\ g\ of\ anhydrous\ Na_2CO_3\ produces\ 491.4\ J\ of\ energy\ \ when\ dissolved\ in\ water$$
$$\therefore 106\ g\ of\ anhydrous\ Na_2CO_3\ produces \frac{106}{2.12} \times 491.4 = 24570\ J = 24.57\ kJ\ of\ energy$$

The heat of solution is –24.57 kJ

Heat of Combustion

When 0.4 g of methanol (CH_3OH) is burnt, the heat evolved produces a 7°C rise in the temperature of 304 g of H_2O. Calculate the molar heat of combustion of methanol. (RAM: C = 12, H = 1, O = 16)

$$Heat\ evolved = mc\Delta\theta = 304\ g \times 4.2\ J\ g^{-1}℃^{-1} \times 7℃ = 8937.6\ J$$
$$0.4\ g\ of\ CH_3OH\ produces\ 8937.6\ J\ of\ heat$$
$$\therefore 32\ g\ of\ CH_3OH\ produces \frac{32}{0.4} \times 8937.6 = 715008\ J\ of\ heat$$

The heat of combustion is –715 kJ