Topic 5: Energetics
5.1 Exothermic and endothermic reactions

- Chemical reactions involve the breaking and making of bonds.

- Breaking bonds requires energy, whereas energy is given out when new bonds are formed.

- Energy: the ability to do work (= move a force through a distance) e.g. heat, light, sound, electricity, chemical energy

\[
\text{Energy (J)} = \text{force (N)} \times \text{distance (m)}
\]
• Heat energy increases the movement of particles in a disordered fashion (=increases the temperature).

• **Heat:** The **total** energy in a given amount of substance *(depends on the amount of substance).*

• **Temperature:** The **average** kinetic energy of the particles in a substance *(is independent on the amount of substance).*

Both beakers have the **same temperature** but the 1 L beaker contains twice as much heat as the 500 L.
• **Enthalpy:** Heat content of a substance “heat inside”.

• It is impossible to measure the actual heat content (the absolute values of the enthalpy) for reactants and products, but the enthalphy change ($\Delta H$) for a reaction can be measured.
An open system can exchange energy and matter with the surrounding, whereas a closed system can exchange only energy.
5.1.1 Exothermic reactions

- Heat is given out from the system to the surrounding during the reaction.
- The heat content of the system decreases, $\Delta H$ is negative ($\Delta H < 0$).
- The products are more stable than the reactants.
Examples of exothermic reactions

- Combustion
- Neutralization
5.1.2. Endothermic reactions

- Heat is absorbed by the system.
- The heat content of the system increases, $\Delta H$ is positive ($\Delta H > 0$).
- The products are less stable than the reactants, but the reaction occur because there is an increase in disorder of the system.
Examples of endothermic processes

- Photosynthesis

- Change of state
  
  \[ s \rightarrow l \]
  
  \[ s \rightarrow g \]
  
  \[ l \rightarrow g \]
5.1.3. Standard enthalphy change of reaction ($\Delta H^\Theta$)

- $\Delta H$ is usually measured under standard conditions:
  
  - a **constant** pressure of 101.3 kPa ($1.013 \times 10^5$ Pa, 1 atmosphere)
  - a temperature of 298K (25ºC)
  - concentrations of 1 mol dm$^{-3}$
  - the most thermodynamically stable allotrope (in case of carbon graphite)
5.2 Calculation of enthalpy changes

- When a substance is heated, its temperature increases.

- The size of the increase depends on:
  - The mass of the substance
  - The heat added
  - The nature of the substance (its heat capacity)

- **Heat capacity**: The amount of heat energy required to raise the temperature of a substance by one kelvin (degree Celsius).
5.2.1 The specific heat capacity (c)

- The amount of heat energy required to raise the temperature of a unit mass (often grams) of a substance by one kelvin (degree Celsius).

- The heat involved in changing the temperature of any substance can be calculated from the equation:

\[ q = m \times c \times \Delta T \]

- The lower the specific heat capacity of a substance, the greater its temperature rise for the same amount of heat absorbed!
• Ex. 1

How much energy is required to increase the temperature of 20 grams of nickel (specific heat capacity 440 J kg\(^{-1}\) °C\(^{-1}\)) from 50°C to 70°C?
Measuring enthalpy changes

• Water is able to absorb large amounts of heat.

• The actual amount of heat evolved in an exothermic reaction can be measured by using the heat given out to the surrounding to increase the temperature of a known mass of water.

• 1 gram of water requires 4.18 J of heat energy for the temperature of the water to increase by 1°C (1 K).

• This is known as the specific heat capacity of water, 4.18 kJ kg$^{-1}$ K$^{-1}$.
Calorimetry

- Two assumptions are made:
  - The reaction is so fast that the maximum temperature is reached before the mixture begins to cool to room temperature.
  - No heat is lost to the surroundings (insulated system)
• To compensate for the heat lost to the surrounding a plot of temperature against time can be drawn.

• By extrapolating the line to the time of mixing, the temperature change that would have taken place (if no heat was lost) can be calculated.

\[ \text{Zn (s) + Cu(SO}_4\text{ (aq) } \rightarrow \text{ZnSO}_4\text{ (aq) + Cu (s)} \]
5.2.2 Heat of combustion

- The heat produced from the combustion of fuels, such as methanol, can be measured by a simple calorimeter.

\[ 2 \text{CH}_3\text{OH (l)} + 3 \text{O}_2 (g) \rightarrow 2 \text{CO}_2 (g) + 4 \text{H}_2\text{O (l)} \]

- The enthalpy change of combustion: The heat produced when one mole of a substance is burned in excess oxygen.

**Ex 2.** When 0.75 g of methanol was burned *completely* it produced a temperature rise of 34.5°C in 100 ml of water. Calculate the \( \Delta H \) for the reaction.
5.2.3 Enthalpy change of neutralization

- The standard enthalpy change of neutralization is the energy change that takes place when one mole of H\(^+\) ions is completely neutralized under standard conditions.

Ex. NaOH (aq) + HCl (aq) → NaCl (aq) + H\(_2\)O (l)

\[ \Delta H^\Theta = -57 \text{ kJ mol}^{-1} \]
• The ΔH of neutralization of a **strong acid with a strong base** is almost the same for all strong acids and strong bases.

• This is because they all dissociate completely in water and the reaction is the same for all, the combination of H\(^+\) and OH\(^-\) ions to form water.

\[
\text{H}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{H}_2\text{O} (\text{l}) \quad \Delta H^\theta = -57 \text{ kJ mol}^{-1}
\]

• For sulfuric acid the equation is:

\[
\frac{1}{2} \text{H}_2\text{SO}_4 (\text{aq}) + \text{KOH} (\text{aq}) \rightarrow \frac{1}{2} \text{K}_2\text{SO}_4 (\text{aq}) + \text{H}_2\text{O} (\text{l})
\]

\[
\Delta H^\theta = -57 \text{ kJ mol}^{-1}
\]
• Ex 3.

50.00 cm$^3$ of 1.0 mol dm$^{-3}$ hydrochloric acid was added to 50.00 cm$^3$ of 1.0 mol dm$^{-3}$ sodium hydroxide solution. The temperature rose by 6.8 °C. Calculate the enthalpy change of neutralization for this reaction. Assume that the density of the solution is 1.00 g cm$^{-3}$ and the specific heat capacity of the solution is 4.18 J g$^{-1}$ ºC$^{-1}$.
5.3 Hess law:

- The overall enthalphy change in a reaction is constant and does not depend on the pathway taken.

- If a reaction consists of a number of steps, the overall enthalpy change is equal to the sum of $\Delta H$ for all the individual steps.

$$\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4$$
• Hess law is a statement of the law of conservation of energy.
• It can be used for calculating the enthalpy change of reactions for which the ΔH cannot be measured directly by experiment.

• **Ex 1** *(Ex 8 p.170 (HL))*

Calculate the $\Delta H^\circ$ for the reaction:

C (s) + $\frac{1}{2}$ O$_2$ (g) $\rightarrow$ CO (g)
• Ex 2. Calculate the $\Delta H^\Theta$ for the conversion of graphite to diamond.

\[
\text{C (s, graphite) + O}_2 (g) \rightarrow \text{CO}_2 (g) \quad \Delta H^\Theta = -393 \text{ kJ mol}^{-1} \\
\text{C (s, diamond) + O}_2 (g) \rightarrow \text{CO}_2 (g) \quad \Delta H^\Theta = -395 \text{ kJ mol}^{-1}
\]

• Energy-cycle-method
• Algebraic method

Reversing the direction of a reaction reverses the sign of $\Delta H$!