5.4 - Bond Enthalpies

5.4.1 - Define the term *average bond enthalpy*

Average Bond Enthalpy - The amount of energy required to break one mole of bonds in the gaseous state averaged across a range of compounds containing that bond.

For example, in determining the bond enthalpy of C-H:

$$CH_{4\,(g)} \rightarrow C_{(g)} + 4H_{(g)} \qquad \Delta H = +1662\ kJ\ mol^{-1}$$

5.4.2 - Explain, in terms of average bond enthalpies, why some reactions are exothermic and others are endothermic

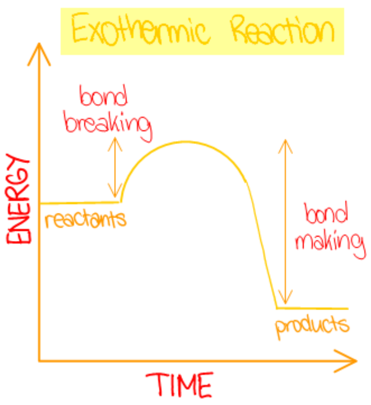
Bond breaking

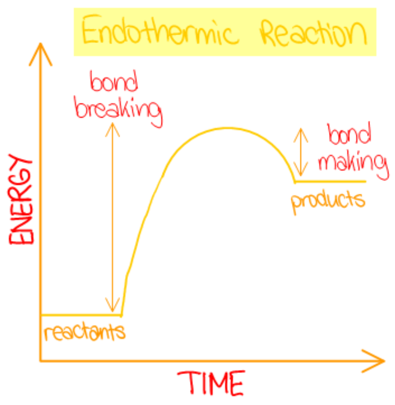
This is an endothermic process, as energy is absorbed to break the bonds between the atoms.

Bond making

This is an exothermic process, as energy is released when the bonds form between the atoms.

If more energy is released in the formation of bonds in the products that was required to break the bonds of the reactants, then the reaction is **exothermic.** If more energy is required to break the bonds of the reactants than is released in the formation of bonds in the products, then the reaction is **endothermic**





The bond enthalpies are used to calculate the enthalpies of the reaction. This is only approximate because only average bond enthalpies are used

$$Enthalpy\ of\ Reaction = \sum D\ (bonds\ broken) - \sum D\ (bonds\ formed$$

∑ is the sum of the terms

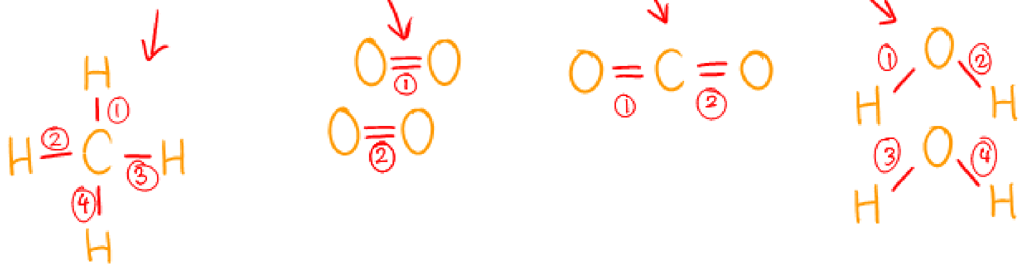
D is the bond enthalpy per mole of bonds

For example:

Combustion of Methane:

$$CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2H_2O_{(g)}$$

$$\Delta H = [4\,D_{C-H} + 2\,D_{O=O}] - [2\,D_{C=O} + 4\,D_{H-O}]$$



$$\Delta H = [4 \times 412 + 2 \times 496] - [2 \times 743 + 4 \times 463]$$

$$= 2640 - 3338$$

$$= -698\ kJ\ mol^{-1}$$