Enthalpy changes: ΔH of reaction, formation, combustion ad neutralisation

Enthalpy, H, is the heat content that is stored in a chemical system. The chemical system is the reactants and products.

The surroundings are what are outside the chemical system.

Heat loss in a chemical system = Heat gain to surroundings (accompanied by a temperature increase) Heat gain in a chemical system = heat loss from surroundings (accompanied by a temperature decrease)

Enthalpy change, ΔH , is the heat exchange with surroundings during a chemical reaction, at constant pressure.

In an exothermic reaction, the enthalpy of the products is smaller than the enthalpy of the reactants. There is heat loss from the chemical system to the surroundings ($\Delta H - ve$). ΔH has a negative sign because heat has been lost by the chemical system.

In endothermic reactions, the enthalpy of the products is greater than the enthalpy of the reactants. There is a heat gain to the chemical system from the surrounding (ΔH + ve). ΔH has a positive sign because heat has been gained by the chemical system.

Activation energy is the minimum energy required to start a reaction by the breaking of bonds.

Standard conditions are a pressure of 100kPa (1 atmosphere), a stated temperature, usually 298K (25°C) and a concentration of 1.0 mol dm⁻³ (for reactions with aqueous solutions).

Standard state is the physical state of a substance under the standard conditions of 100kPa (1 atmosphere) and 298K (25°C).

Enthalpy change of reaction is the enthalpy change associated with a state equation under standard conditions.

The standard enthalpy change of formation is the formation of 1 mol of a compound from its elements under standard conditions.

The standard enthalpy change of combustion is the complete combustion of 1 mol consistence under standard conditions.

The standard enthalpy change of neutralisation is the formation of water from neutralisation under standard conditions.

Standard enthalpy changes determined experimentally can be less/mer exothermic or endothermic than the calculated theoretical values because their may have been measured to the surroundings/absorbed from the surroundings, incomplete reactions in a may occurred (reactions out completely reacted) or it may not have occurred under standard conditions.

Bond enthalpies

Average bond enthalpy is the breaking of 1 mole of bonds in gaseous molecules.

Every species must be a gas. If changing state symbols, you have to consider the enthalpy change of converting between the standard states.

Bond enthalpy is an endothermic change with bonds being broken. When the same bonds are made, the enthalpy change will be the same magnitude but will be exothermic.

Bond breaking absorbs energy and bond making releases energy. In exothermic reactions, more energy is released than absorbed (bonds formed are stronger). In endothermic reactions, more energy is absorbed than released (bonds formed are weaker).

An actual bond enthalpy may differ from the average value because bonds have different strengths in different environments and the average bond enthalpy is an average over many compounds containing the bond.

Simple collision theory

Increasing the concentration increases the rate of reaction because there are more particles per (unit)/ in the same volume so they are closer together and have a greater chance of colliding. There will be more frequent collisions per second.

For a particular length of time, more frequent collisions will take place with greater energy than the activation energy.

Increasing the gas pressure increases the rate of reaction because the same number of molecules occupy a smaller volume (same as concentration).