Energetics / Thermochemistry

Heat is a form of energy that flows from something at a higher temperature to something at a lower temperature.

Temperature is a measure of the average kinetic energy of particles.

In an exothermic reaction, heat energy is transferred from a system to the surroundings - the surrounding gets hotter.

In an endothermic reaction, a system takes in heat energy from the surrounding - the surroundings get cooler.

Enthalpy change is the amount of heat energy taken in/ given out in a chemical reaction.

ΔH for an exothermic ~~is~~ reaction is negative
ΔH for an endothermic reaction is positive

$$\Delta H = \Delta H_{products} - \Delta H_{reactant}$$

Total energy is conserved in a chemical reaction

The term stability is usually used to describe the relative energies of reactants and products in a chemical reaction. If products have less energy than the reactants then they are more stable.

Standard enthalpy change of reaction ($\Delta H_r^{\ominus}$) is the enthalpy change when molar amounts of reactants react together under standard conditions to give products.

pressure = 1.0×10^5 Pa , temperature = 298 K

Standard enthalpy change of combustion ($\Delta H_c^\ominus$) is the enthalpy change when one mole of a substance is completely burnt in oxygen under standard conditions

The energy required to raise the temperature of 1 g of substance by 1 K

$$Q = mc\Delta T$$

Enthalpy change of neutralisation (ΔH_n) is the enthalpy change when one mole of water molecules are formed when an acid (H^+) reacts with an alkali (OH^-) under standard conditions. (always exothermic)

Enthalpy change of solution (ΔH_{sol}) is the entahhalpy change when one mole of solute is dissolved in excess solvent to form a solution of 'infinite' dilution' under standard conditions

- Hess's law

The enthalpy change accompanying a chemical reaction is independent of the pathway between the initial and final states.

Standard enthalpy change of formation ($\Delta H_f^\ominus$) is the enthalpy change when one mole of a substance is formed from its elements in their standard states under standard conditions.

$\Delta H_f^\ominus$ for any element in its standard state is zero

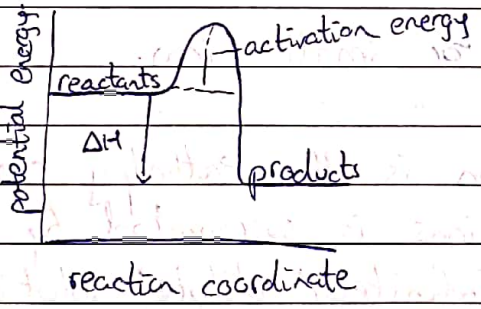
$$\Delta H_r = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$$

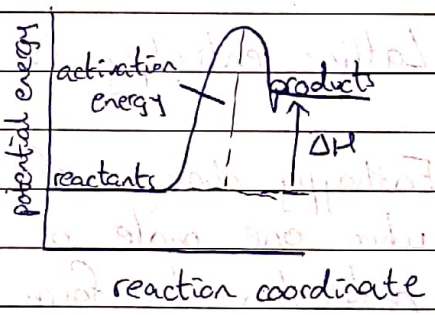
Bond enthalpy is the enthalpy change when one mole of covalent bonds, in a gaseous molecule, is broken under standard conditions.

Bond breaking requires energy (endothermic): ΔH positive
Bond making releases energy (exothermic): ΔH negative

$$\Delta H_B = \Delta H \text{ Bonds broken} - \Delta H \text{ Bonds made}$$

Only gives accurate answer if all the reactants and products are in gas phase and not in solid or liquid state.





Standard enthalpy change of atomisation ($\Delta H^{\ominus}_{at}$) - this is the enthalpy change when one mole of gaseous atoms is formed from an element under standard conditions.

First ionisation energy is the enthalpy change when one electron is removed from each atom in one mole of gaseous atoms under standard conditions:

$$M_{(g)} \longrightarrow M^+ + e^-$$

(all energies are endothermic)

First electron affinity is the enthalpy change when one electron is added to each atom in one mole of gaseous atoms under standard conditions:

$$X_{(g)} + e^- \longrightarrow X^-_{(g)}$$

Lattice enthalpy ($\Delta H^{\ominus}_{latt}$) is the enthalpy change when one mole of an ionic compound is broken apart into its constituent gaseous ions under standard conditions:

Greater the electrostatic attraction between the ions, the more energy has to be supplied to break the lattice apart.

The higher the charges on the ions, the more strongly the will attract each other. ∴ greater lattice enthalpy.

Lattice enthalpy is larger for smaller ions.

Enthalpy change of hydration is the enthalpy change when one mole of gaseous ions is surrounded by water molecules to form an 'infinitely dilute solution' under standard conditions.

Higher the charge more exothermic enthalpy change of hydration. Smaller the ions more exothermic enthalpy change is.

Entropy is a measure of how the available energy is distributed among the particles.

Solids have lower entropy than gas as less freedom of movement and there are fewer ways of distributing the energy among particles.

increasing entropy →

solid liquid gas

An increase in number of moles of gas: ΔS is positive
An decrease in number of moles of gas: ΔS is negative

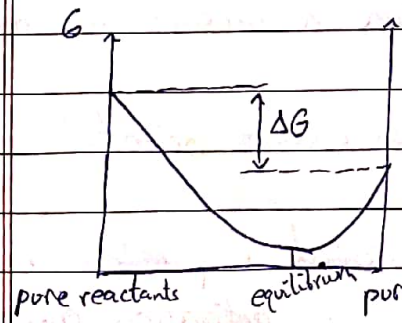
$$\Delta S^{\ominus} = \sum S^{\ominus}_{products} - \sum S^{\ominus}_{reactants}$$

Spontaneous reaction: one that occurs without any outside influence

$$\Delta G = \Delta H - T\Delta S$$

For a reaction to be spontaneous, ΔG, for the reaction must be negative.

ΔH	ΔS	$-T\Delta S$	ΔG	Spontaneous
negative	positive	negative	negative	at all temperature
positive	negative	positive	becomes more negative	(spontaneous) at high temp.
negative	positive	negative	becomes more positive	(not spontaneous) as ↑T
positive	negative	positive	positive	not never



$-\Delta G$ closer to products
$+\Delta G$ closer to reactants