3.1.4 Energetics

3.1.4.1 *Enthalpy change (ΔH). Standard conditions (100 kPa, stated T). Define ΔcH■ and ΔfH■.*

Enthalpy change (ΔH): heat energy change at constant pressure.

Exothermic: ΔH negative. Endothermic: ΔH positive.

Standard conditions: 100 kPa, usually 298 K.

ΔcH■: enthalpy change when one mole of a substance is completely burned in excess oxygen under standard conditions.

ΔfH■: enthalpy change when one mole of a compound is formed from its elements in their standard states. ΔfH■ of an element in its standard state = 0.

3.1.4.2 *Calorimetry: q = mcΔT. Calculate molar enthalpy changes from experimental results.*

q = mcΔT where q = energy (J), m = mass of solution (g), c = 4.18 J g■¹ K■¹, ΔT = temperature change.

ΔH = –q / n (kJ mol■¹). Negative sign: if solution heats up, reaction is exothermic.

Assumptions: no heat loss; solution has density and heat capacity of water; all solute reacts.

3.1.4.3 *Hess's Law: calculate enthalpy changes from enthalpies of formation or combustion.*

Hess's Law: total enthalpy change is independent of the route taken.

Using ΔfH■: ΔrH■ = ΣΔfH■(products) – ΣΔfH■(reactants)

Using ΔcH■: ΔrH■ = ΣΔcH■(reactants) – ΣΔcH■(products)

Draw an enthalpy cycle showing two routes from reactants to products.

3.1.4.4 *Mean bond enthalpy. Calculate approximate ΔH for gaseous reactions. Explain why values differ from Hess's Law.*

Mean bond enthalpy: average energy to break one mole of a bond in the gaseous state, averaged across different molecules.

ΔH = Σ(bonds broken) – Σ(bonds formed)

Bonds broken = endothermic (+). Bonds formed = exothermic (–).

Values differ from Hess's Law: mean bond enthalpies are averages — actual values depend on molecular environment. Only valid for gaseous reactions.

★ REQUIRED PRACTICAL RP2

Measurement of an Enthalpy Change

Method (neutralisation/displacement/dissolution): measure known volume of solution into polystyrene cup. Record baseline temperature for 2–3 min. Add second reagent at a known time. Stir and record temperature every minute for 5–6 min. Plot T vs time. Extrapolate to time of mixing to find true ΔT. Calculate $q = mc\Delta T$ then $\Delta H = -q/n$.

For combustion: weigh spirit burner, burn under copper calorimeter with known water, record ΔT, reweigh.

Uncertainties: thermometer ±0.5°C; measuring cylinder ±0.5 cm^3; balance ±0.001 g.

Common errors: heat loss to surroundings (ΔH less exothermic than true); assuming solution density = water; not using excess reagent; for combustion: heat loss to air, incomplete combustion.

Improvements: polystyrene cup with lid; T vs time extrapolation; digital thermometer; draught shield for combustion.