# Topic 8: Energetics I

Edexcel A Level Chemistry • Paper 1 (9CH0/01)

**8.1** Know that standard conditions are 100 kPa and a specified temperature, usually 298 K.

Standard conditions: pressure = 100 kPa, temperature = 298 K (25°C). Standard state is the physical state of a substance under these conditions. The standard state symbol is ⦵ (or °). All standard enthalpy changes are measured under these conditions.

**8.2** Know that the enthalpy change is the heat energy change measured at constant pressure.

Enthalpy change (ΔH) is the heat energy transferred in a reaction at constant pressure. If heat is given out to the surroundings, the reaction is exothermic (ΔH is negative). If heat is absorbed from the surroundings, the reaction is endothermic (ΔH is positive).

**8.3** Be able to construct and interpret enthalpy level diagrams showing an enthalpy change, including appropriate signs for exothermic and endothermic reactions.

Enthalpy level diagrams show reactants and products on an enthalpy axis.

Exothermic: products are lower than reactants; the arrow points downward; ΔH is negative.

Endothermic: products are higher than reactants; the arrow points upward; ΔH is positive.

Note: activation energy is NOT shown on enthalpy level diagrams (it is shown on reaction profile diagrams).

**8.4** Be able to define standard enthalpy change of: (i) reaction, (ii) formation, (iii) combustion, (iv) neutralisation.

(i) Standard enthalpy of reaction ($\Delta H^{\ominus}_r$): the enthalpy change when the amounts shown in the equation react under standard conditions.

(ii) Standard enthalpy of formation ($\Delta H^{\ominus}_f$): the enthalpy change when one mole of a compound is formed from its elements in their standard states under standard conditions. By definition, $\Delta H^{\ominus}_f$ of an element in its standard state = 0.

(iii) Standard enthalpy of combustion ($\Delta H^{\ominus}_c$): the enthalpy change when one mole of a substance is completely burned in excess oxygen under standard conditions. Always exothermic (negative).

(iv) Standard enthalpy of neutralisation ($\Delta H^{\ominus}_{neut}$): the enthalpy change when an acid and base react to form one mole of water under standard conditions.

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8.5 Understand experiments to measure enthalpy changes in terms of: (i) processing results using Q = mcΔT, (ii) evaluating sources of error and assumptions made.

Q = mcΔT where Q = energy transferred (J), m = mass of solution (g), c = specific heat capacity (4.18 J $g^{-1}$ $K^{-1}$ for water), ΔT = temperature change (K or °C).

Then: ΔH = –Q / n (in kJ $mol^{-1}$), where n = moles of the limiting reagent. The negative sign accounts for the direction of energy transfer.

Key assumptions: all heat is transferred to the solution; the solution has the density and specific heat capacity of water; no heat is lost to the surroundings.

Types of experiment: mixing in an insulated container (e.g. polystyrene cup), spirit burner combustion, or indirect measurement via Hess's Law.

8.6 Be able to calculate enthalpy changes in kJ $mol^{-1}$ from given experimental results.

Method: (1) Calculate Q = mcΔT in joules. (2) Convert to kJ by dividing by 1000. (3) Calculate moles of the substance of interest. (4) Divide Q (kJ) by moles to get kJ $mol^{-1}$. (5) Assign the correct sign (– for exothermic, + for endothermic). Always give both a sign and units in the final answer.

8.7 Be able to construct enthalpy cycles using Hess's Law.

Hess's Law: the total enthalpy change for a reaction is independent of the route taken, provided the initial and final conditions are the same. This allows calculation of enthalpy changes that cannot be measured directly.

Draw a cycle with two routes from reactants to products:

Route 1: direct (the unknown ΔH)

Route 2: indirect, via intermediate compounds (using known ΔH values)

By Hess's Law: ΔH (Route 1) = ΔH (Route 2). Common cycles use enthalpies of formation or enthalpies of combustion as the intermediate step.

8.8 Be able to calculate enthalpy changes from data using Hess's Law.

Using enthalpies of formation:

ΔHr = ΣΔHf (products) – ΣΔHf (reactants)

Using enthalpies of combustion:

ΔHr = ΣΔHc (reactants) – ΣΔHc (products)

Remember: multiply ΔH values by the stoichiometric coefficients.

ΔHf of elements in their standard states = 0.

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8.9 Know what is meant by the terms 'bond enthalpy' and 'mean bond enthalpy'.

Bond enthalpy: the energy required to break one mole of a specific covalent bond in a gaseous molecule, with all species in the gaseous state. It is always endothermic (positive).

Mean bond enthalpy: the average energy required to break one mole of a particular type of bond across a range of different gaseous molecules. It is a mean value because the exact bond enthalpy varies depending on the molecular environment.

8.10 Be able to calculate an enthalpy change of reaction using mean bond enthalpies and explain the limitations of this method.

$\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$

Bonds broken = energy in (endothermic, positive). Bonds formed = energy out (exothermic, negative).

Limitations:

- Mean bond enthalpies are averages – the actual bond enthalpy in a specific molecule may differ, so the calculated ΔH is only an estimate.
- Only applies to reactions where all reactants and products are gaseous (otherwise you would need to account for enthalpy changes of vaporisation/fusion).
- Does not account for intermolecular forces.

8.11 Be able to calculate mean bond enthalpies from enthalpy changes of reaction.

Rearrange $\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$ to solve for the unknown bond enthalpy. Substitute all known bond enthalpies and the overall ΔH, then solve algebraically. You may need to account for multiple bonds of the same type (e.g. 4 × C–H in $CH_4$).

## To Determine the Enthalpy Change of a Reaction Using Hess's Law

ABOUT THESE NOTES

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