

AQA Chemistry A-level

3.1.4: Energetics

Detailed Notes

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3.1.4.1 - Enthalpy Change (ΔH°)

In a reaction, bonds are broken and then made. For bonds to be **broken**, energy is **taken in** from the surroundings and when bonds are **formed**, energy is **given out**. The overall energy change of the reaction depends on how much energy is transferred in these processes.

When energy is taken in from the surroundings, the **enthalpy change is positive**. When energy is released, the **enthalpy change is negative**.

Overall enthalpy change (ΔH) can be calculated:

$$\Delta H = \text{energy to break bonds (+ve)} + \text{energy to make bonds (-ve)}$$

Measuring Enthalpy Change

Enthalpy is measured under standard conditions of **100 kPa pressure** and a temperature of **298 K**.

Enthalpy of Formation ($\Delta_f H^\circ$)

This is defined as:

The enthalpy change when one mole of a substance is produced from its elements under standard conditions.

Enthalpy of Combustion ($\Delta_c H^\circ$)

This is defined as:

The enthalpy change when one mole of a substance is burned completely in oxygen under standard conditions.

Mean Bond Enthalpies

Different covalent bonds require different amounts of energy to be broken. Values can be found experimentally using **calorimetry** methods. The bond enthalpy values calculated in this way often differ from the **data book values** as they are **not exact** and **vary in each situation**. This means the data book values are **averaged values**.

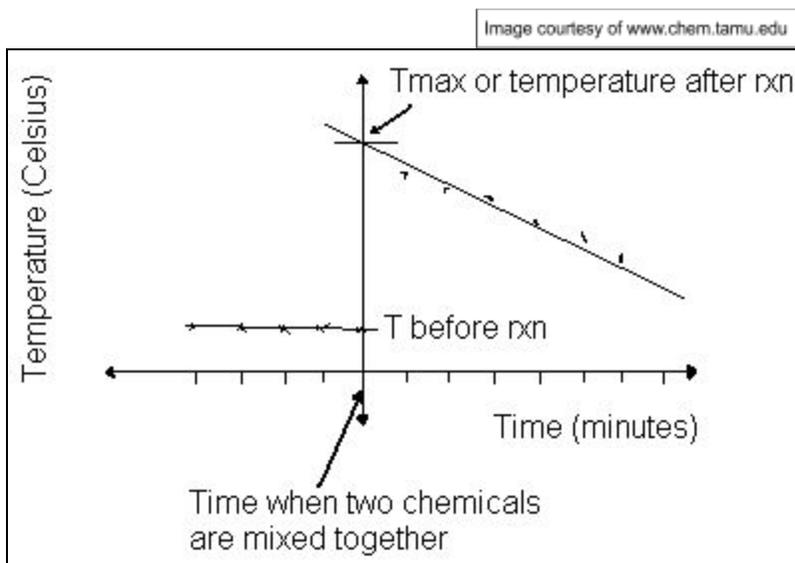




3.1.4.2 - Calorimetry

Calorimetry is an **experimental method** for finding enthalpy change by measuring **temperature change over time**. When observed and plotted on a graph, data can be **extrapolated** to give an accurate value for the change in temperature.

Example:



This measured change in temperature is **proportional** to the energy change:

$$q = mc\Delta T$$

(q = energy change) (m = mass) (c = specific heat capacity) (ΔT = temperature change)

It allows the energy in or out of a substance of known mass to be calculated, which depends on its **specific heat capacity**, defined as:

The energy required to raise 1g of substance by 1K without a change of state.

Using this energy value (q), **enthalpy change per mole** of substance can be calculated:

$$\Delta H = \frac{q}{\text{moles}}$$

ΔH values found using calorimetry are **never completely accurate** as energy is easily lost from the system through **conduction** or **convection** and inaccuracies in measuring temperatures.



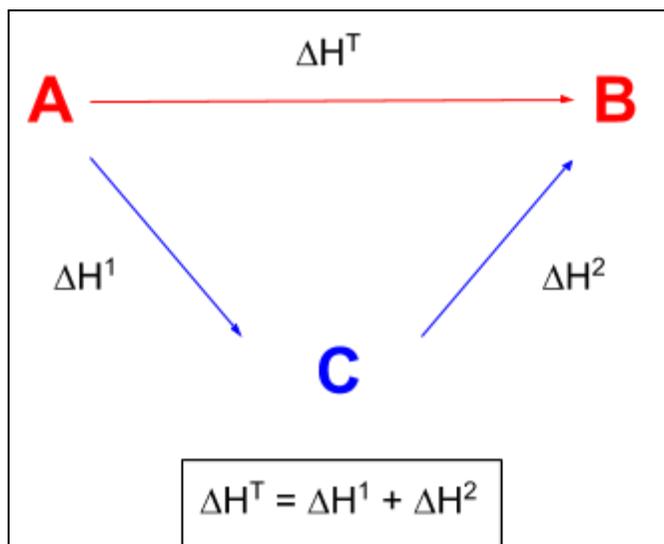


3.1.4.3 - Hess's Law

Energy in a reaction system must be **conserved**, as it cannot be created or destroyed. Therefore the **overall enthalpy change** for a reaction is **the same**, regardless of the **route taken**.

This is **Hess's Law** and it is used to determine the enthalpy changes for reactions that cannot be found directly using an experimental method. To do this, a **triangular cycle method** is used with an inbetween product. The direction of the arrows indicates whether the values should be added or taken away. (They can be treated like vectors):

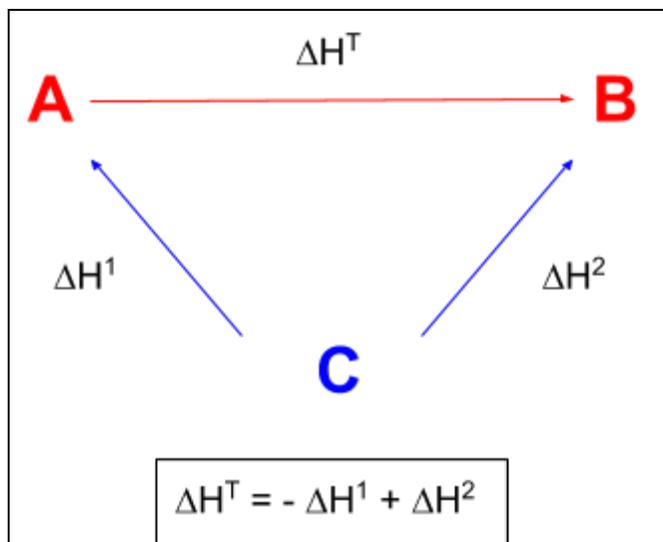
Example:



Enthalpies of Formation

Hess's law can be used to calculate an enthalpy change using given enthalpies of formation. When setting up the triangular diagram, the **arrows point out from the central product C** as both A and B are formed from the elements at C.

Example:



Note how ΔH^1 is subtracted as the reaction goes in the opposite direction to the arrow.

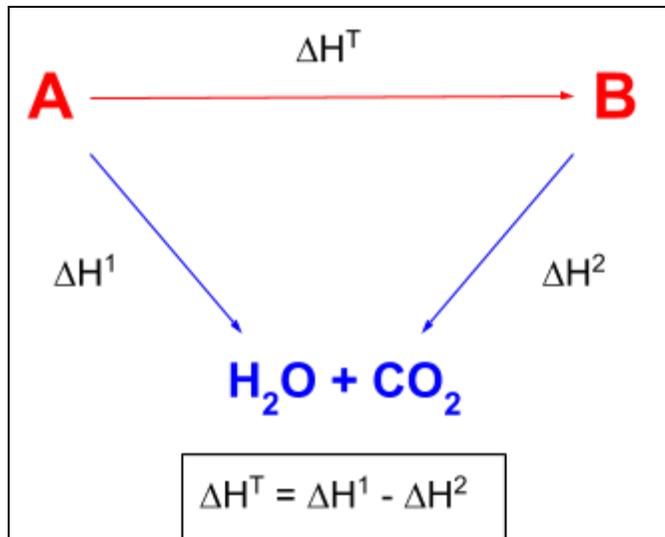




Enthalpies of Combustion

Hess's law can also be used to calculate an enthalpy change using given enthalpies of combustion. When setting up the triangular diagram, the **arrows point towards the central product (which is always H₂O and CO₂)** as both A and B burn to form the products at C.

Example:



Note how ΔH^2 is subtracted as the reaction goes in the opposite direction to the arrow.

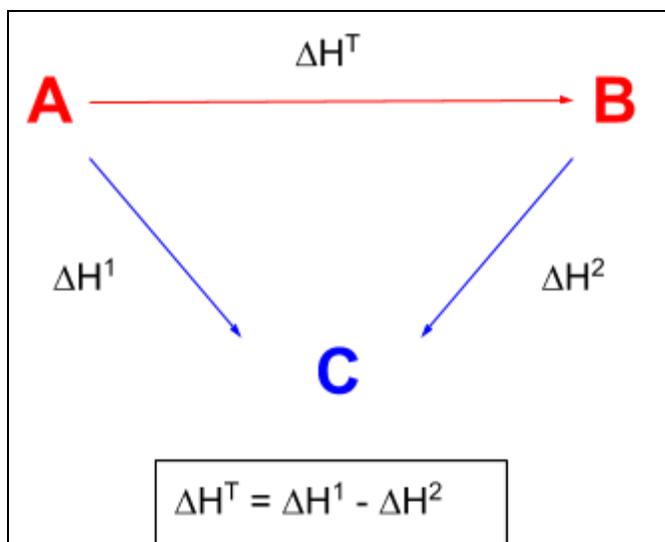
3.1.4.4 - Bond Enthalpies

Bond enthalpy data is an **averaged** value representing:

The energy required to break one mole of the stated bond in a gaseous state, under standard conditions.

They can be used alongside Hess's Law to find the **overall enthalpy change** for a reaction. The reaction **arrows always point towards the central product**.

Example:



Note how ΔH^2 is subtracted as the reaction goes in the opposite direction to the arrow.

