

Chemical energetics

A-Level Chemistry 9701 · Topic 5 revision notes

Everything from the 4 lessons of this topic on a few pages: the key idea of each lesson, the diagram to remember, the points to learn and the traps examiners set. Read it, cover it, test yourself.

1. Enthalpy changes and definitions

3. Bond energies

2. Calorimetry

4. Hess's law

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1 Enthalpy changes and definitions

Syllabus 5.1

An enthalpy change, ΔH , is the heat energy exchanged with the surroundings at constant pressure. **Exothermic** reactions release heat, so ΔH is **negative**; **endothermic** reactions absorb heat, so ΔH is **positive**. Standard values ($\ominus$) are for 298 K, 101 kPa and 1 mol dm^{-3} solutions, with substances in their standard states. Learn four definitions exactly: reaction (the molar quantities in the equation), **formation** (1 mol of compound from its elements), **combustion** (1 mol burned completely in oxygen) and **neutralisation** (1 mol of water formed).

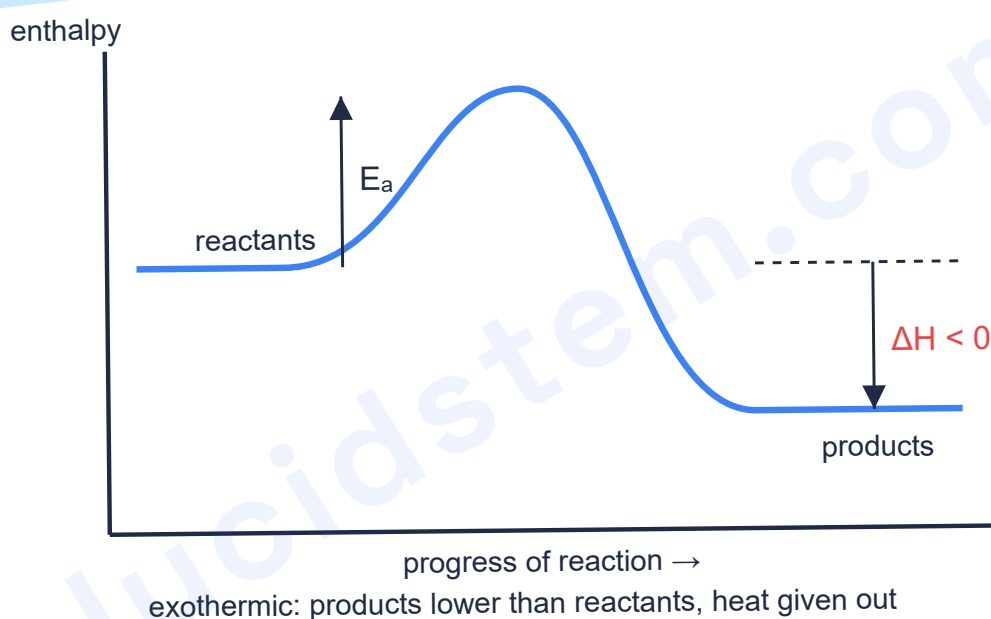


Fig. 1: reaction pathway diagram for an exothermic reaction

One mole, standard states, right sign.

- ✓ **Exothermic:** ΔH negative;
- endothermic:** positive.
- ✓ **Standard:** 298 K, 101 kPa, 1 mol dm^{-3} .
- ✓ **Formation:** 1 mol of compound from elements; elements are 0.
- ✓ **Combustion:** 1 mol burned completely;
- neutralisation:** 1 mol of water.

Watch out!

The standard enthalpy change of formation of water and the standard enthalpy change of combustion of hydrogen are the **same** equation, $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$, so they have the same value, -286 kJ mol^{-1} . Writing $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ doubles the value and fits neither definition.

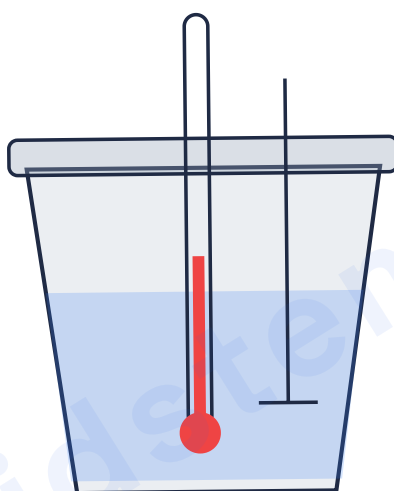
2 Calorimetry

Syllabus 5.1

The heat transferred to or from the surroundings is $q = mc\Delta T$, where m is the mass of water or solution heated (not the mass of reactant), $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$ and ΔT is the temperature change. Divide by the amount in moles of the reactant that was used up (the limiting one) and convert J to kJ: $\Delta H = -q \div n$. The minus sign means a temperature rise gives a **negative** (exothermic) ΔH . Heat loss to the surroundings makes experimental values less exothermic than data book values, most of all in combustion experiments.

thermometer
 $\Delta T = T_2 - T_1$

stirrer



lid
reduces heat loss

polystyrene cup
insulates

solution
mass m , heat capacity c

$$q = mc\Delta T, \text{ then } \Delta H = -q \div n$$

Fig. 1: a simple polystyrene cup calorimeter for reactions in solution

$$q = mc\Delta T, \text{ then } \Delta H = -q \div n.$$

- ✓ m is the mass of water or solution ($1 \text{ cm}^3 = 1 \text{ g}$).
- ✓ $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$; q in J, divide by 1000.
- ✓ n is the limiting reagent (or water formed).
- ✓ **Temperature rise** means ΔH negative.

Watch out!

In $q = mc\Delta T$, m is the mass of the water or solution whose temperature changes, never the mass of fuel or solid added.

Remember too that q comes out in joules: divide by 1000 for kJ, and add the minus sign yourself when the temperature rose.

3 Bond energies

Syllabus 5.1

Bond energy is the energy needed to break one mole of a particular covalent bond in **gaseous** molecules. Breaking bonds is endothermic and forming bonds is exothermic, so $\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$. For bonds such as C-H the value depends on the molecule, so data books quote a **mean** bond energy averaged over many compounds. That, and substances that are not gases under standard conditions, is why bond energy answers differ slightly from experimental ΔH values.

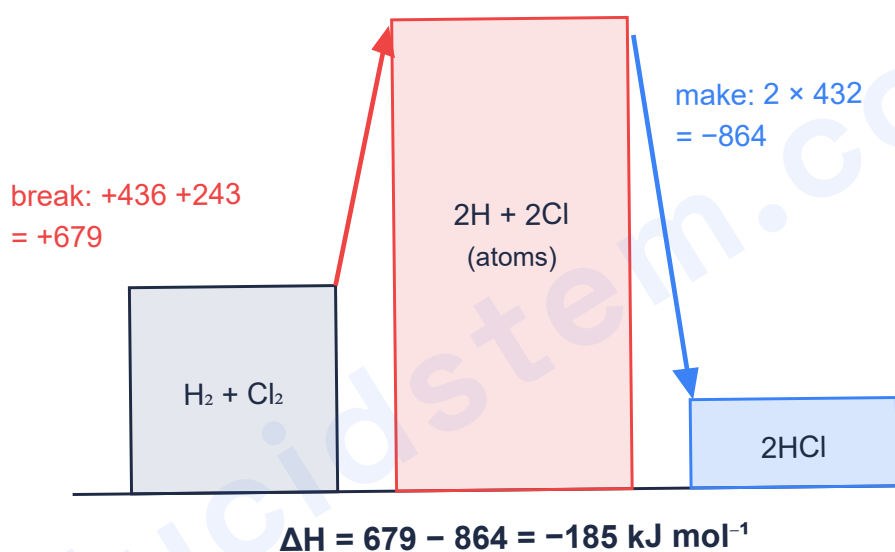


Fig. 1: breaking the bonds in H_2 and Cl_2 , then forming two $\text{H}-\text{Cl}$ bonds

$\Delta H = \text{broken} - \text{formed}$.

- ✓ **Draw** displayed formulae; count every bond.
- ✓ **Multiply** by the coefficients in the equation.
- ✓ **Mean** bond energies are averages over many compounds.
- ✓ **Gas phase only:** liquids cause discrepancies.

Watch out!

It is **broken minus formed**, not the other way round; reversing it flips the sign. And count bonds per molecule and per mole of equation: $2\text{H}_2\text{O}$ contains four O-H bonds, and CO_2 contains two C=O double bonds, not two single C-O bonds.

4 Hess's law

Syllabus 5.2

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. Build a cycle with the unknown change along the top and a route through measurable changes. With **formation** data the common point is the elements below: $\Delta H_r = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$. With **combustion** data the common point is the combustion products below: $\Delta H_r = \sum \Delta H_c(\text{reactants}) - \sum \Delta H_c(\text{products})$. Multiply each value by its coefficient and reverse the sign of any arrow you travel against.

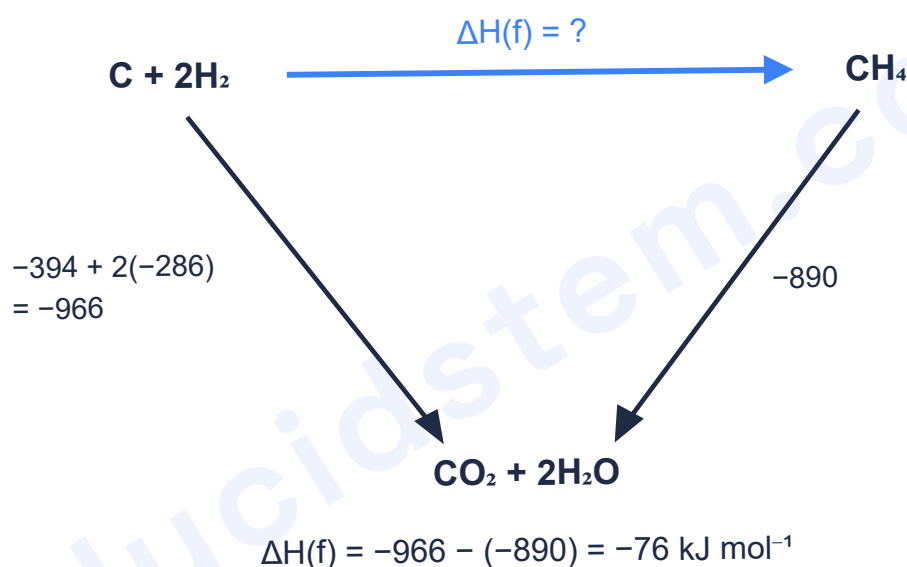


Fig. 1: a Hess cycle giving the enthalpy change of formation of methane from combustion data

Draw the cycle, then follow the arrows.

- ✓ **Formation:** products – reactants.
- ✓ **Combustion:** reactants – products.
- ✓ **Multiply** by coefficients; elements have $\Delta H_f = 0$.
- ✓ **Against an arrow:** change the sign.

Watch out!

The two formula shortcuts have opposite orders: formation is **products minus reactants**, combustion is **reactants minus products**. If unsure, draw the cycle and follow the arrows, changing the sign whenever you travel against one. Never forget to multiply by the coefficients, and remember that elements have $\Delta H_f = 0$.

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