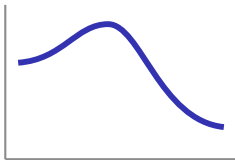


Chemical Energetics

A-Level Chemistry ■ Topic 5

A-Level Chemistry



What we will cover

- 1 Enthalpy change ΔH
- 2 Types of enthalpy change
- 3 Bond energy
- 4 Measuring ΔH
- 5 Hess's law



burning fuel is exothermic



Enthalpy change ΔH

Enthalpy change ΔH is the heat transferred at constant pressure.

exothermic

$\Delta H < 0$ – heat given
out; products more stable

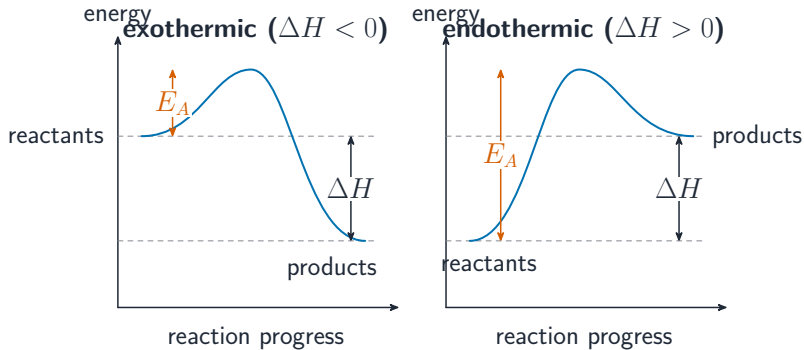
endothermic

$\Delta H > 0$ – heat
taken *in*; a cold pack

Reaction pathway diagrams

- exothermic: products sit **lower** than reactants ($\Delta H < 0$).
- endothermic: products sit **higher** than reactants ($\Delta H > 0$).

Reaction pathway diagrams, in detail



endothermic: products sit **higher** than reactants ($\Delta H > 0$).

Types of enthalpy change

$\Delta H_r^\ominus$ **reaction** – for the equation as written

$\Delta H_c^\ominus$ **combustion** – 1 mol burnt in oxygen

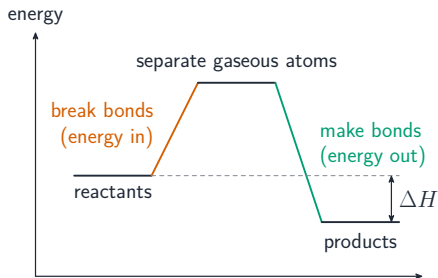
$\Delta H_f^\ominus$ **formation** – 1 mol of compound from elements

$\Delta H_{\text{neut}}^\ominus$ **neutralisation** – 1 mol of water formed

Standard conditions ($^\ominus$): 298 K, 101 kPa, each substance in its normal state.



Bond energy



Breaking bonds takes energy in;
making bonds gives it out:

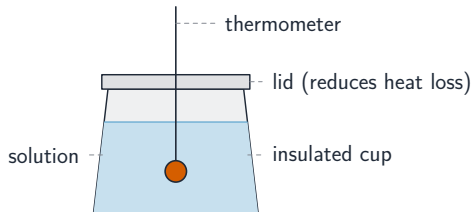
$$\Delta H_r = \sum(\text{broken}) - \sum(\text{made})$$



$$(436 + 242) - (2 \times 431) = -184 \text{ kJ mol}^{-1}$$

Breaking bonds takes energy in; making bonds gives energy out. ΔH is the difference between the two

Measuring ΔH



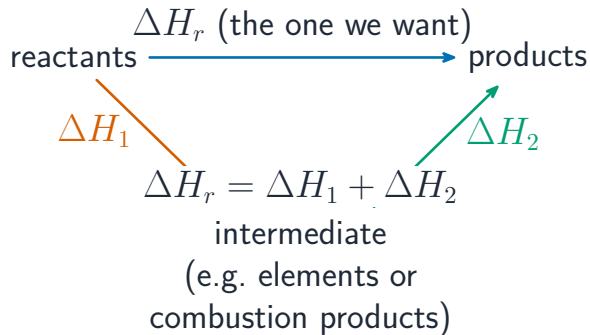
Heat transferred: $q = mc\Delta T$ (**specific heat capacity** c).

$$\Delta H = -\frac{mc\Delta T}{n}$$

0.50 g methanol heats 100 g water by 18 °C

$$q = 7520 \text{ J}, n = 0.0156 \Rightarrow \Delta H_c \approx -480 \text{ kJ mol}^{-1}$$

Hess's law



Hess's law : the total enthalpy change is the same by any route (energy is conserved).

An **energy cycle** finds a ΔH you cannot measure directly: $\Delta H_r = \Delta H_1 + \Delta H_2$.

Enthalpy change, ΔH , continued



An instant cold pack uses an endothermic reaction that takes in heat.

Standard conditions and types of enthalpy change

formation ΔH_f :

elements



1 mol of the compound

combustion ΔH_c :

1 mol substance + O₂



CO₂ + H₂O

Don't confuse them: formation builds 1 mol of a compound from its elements; combustion burns 1 mol of a substance in oxygen

The named enthalpy changes, and Hess's law

a reaction pathway diagram	shows the energy of reactants and products, and the energy hill between them – the activation energy
of formation	one mole of a compound from its elements in their standard states
of combustion	one mole of a substance burned completely in oxygen
of neutralisation	one mole of water formed from acid and alkali
bond energy	the energy to break one mole of a particular bond in the gas state – so it is always positive
Hess's law	the enthalpy change is the same by any route – so an unmeasurable ΔH can be got from measurable ones

Hess's law is used two ways: from enthalpies of **formation** (products minus reactants) and from enthalpies of **combustion** (reactants minus products). The two subtractions run opposite ways.

The four named enthalpy changes

Reaction

$\Delta H_r^\ominus$, the **enthalpy change of reaction**: for the amounts *shown in the equation*.

Formation

$\Delta H_f^\ominus$, the **enthalpy change of formation**: one mole of a compound forms from its elements in their standard states.

Combustion

$\Delta H_c^\ominus$, the **enthalpy change of combustion**: one mole of a substance burns *completely* in oxygen.

Neutralisation

$\Delta H_{\text{neut}}^\ominus$, the **enthalpy change of neutralisation**: one mole of water forms from an acid and an alkali.

Each definition pins **one mole** of a different thing – the compound, the fuel, the water. Naming the wrong one is the commonest lost mark here.

Exam tips

- Define each enthalpy change with its **exact standard conditions** (e.g. combustion = one mole burned completely in excess oxygen).
- Use $\Delta H = \sum(\text{bonds broken}) - \sum(\text{bonds made})$; breaking is endothermic (+), making is exothermic (−) – getting the sign the wrong way round is the classic error.
- In $q = mc\Delta T$ use the mass of the **water/solution**, then divide by moles and add the minus sign for an exothermic reaction.
- Draw **Hess cycles** with arrows the same way round, follow the alternative route, and always give ΔH a sign and units (kJ mol^{-1}).



Topic 5 – key takeaways

1

Enthalpy

exothermic $\Delta H < 0$; endothermic $\Delta H > 0$

2

Bond energy

$\Delta H = \text{bonds broken} - \text{bonds made}$

3

Calorimetry

$q = mc\Delta T$; $\Delta H = -q/n$

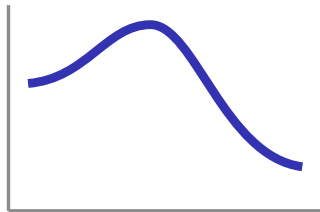
4

Hess's law

ΔH is route-independent

Check yourself

- 1 Sign of ΔH for an exothermic reaction?
- 2 Does breaking a bond take in or give out energy?
- 3 Write the equation for heat transferred.
- 4 What does Hess's law say about the route?



Answers 1. negative 2. takes in energy 3. $q = mc\Delta T$ 4. it is the same by any route