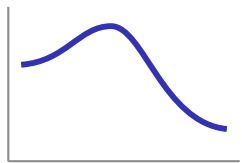
Chemical Energetics

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

1 Enthalpy

Enthalpy change ΔH

Enthalpy change 焓变 ΔH is the heat transferred at constant pressure.

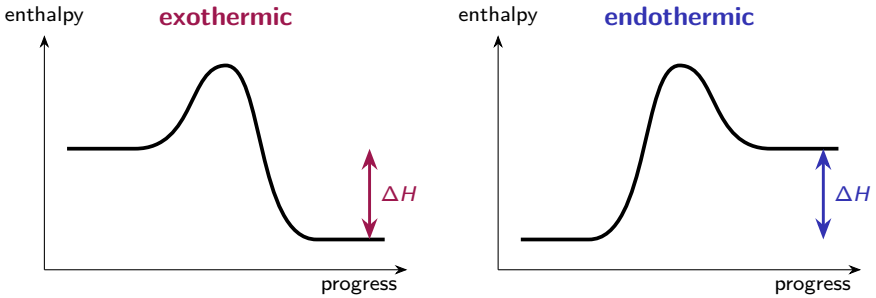
exothermic

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

endothermic

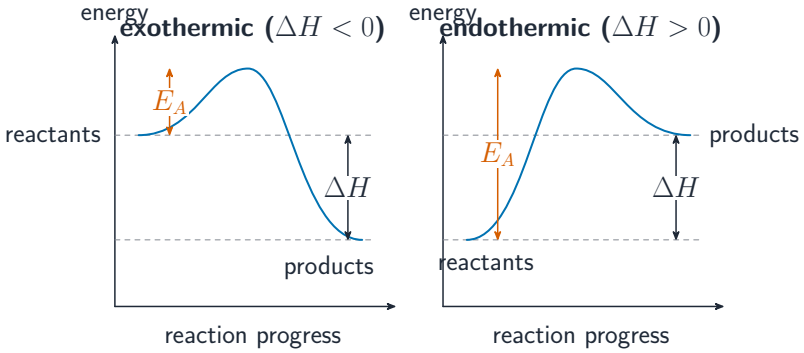
$\Delta H > 0$ – heat taken *in*; a cold pack

Reaction pathway diagrams



- **Exothermic**: products **lower** than reactants, $\Delta H < 0$
- **Endothermic**: products **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 反应焓变** – for the equation as written

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

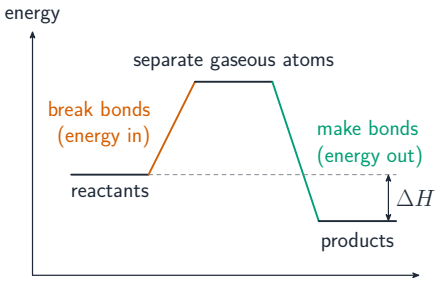
$\Delta H_f^\ominus$ **formation 生成焓变** – 1 mol of compound from elements

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

Standard conditions 标准条件 ($^\ominus$): $298\,\mathrm{K}$, $101\,\mathrm{kPa}$, each substance in its normal state.

2 Calculating ΔH

Bond energy



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

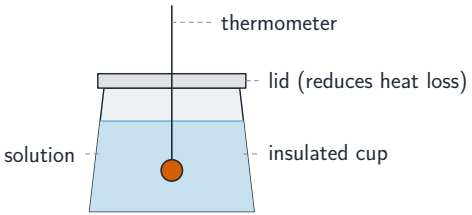
$$\Delta H_r = \sum(\text{broken}) - \sum(\text{made})$$

$H_2 + Cl_2 \rightarrow 2HCl$

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

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

Measuring ΔH



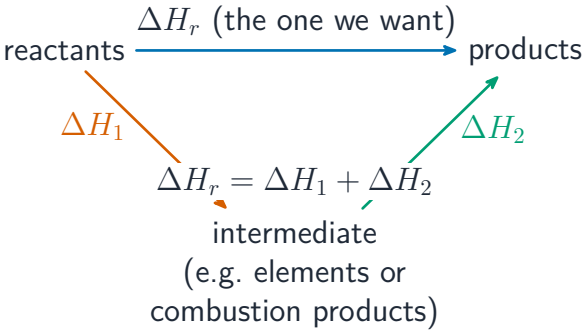
Heat transferred: $q = mc\Delta T$ (**specific heat capacity** 比热容 c).

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

0.50 g methanol heats 100 g water by 18 °C

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

Hess's law



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

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

An endothermic reaction in a cold pack



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 + O_2 → $CO_2 + H_2O$

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 反应路径图	shows the energy of reactants and products, and the energy hill between them – the **activation energy** 活化能
of formation 生成焓变	one mole of a compound from its elements in their standard states
of combustion 燃烧焓变	one mole of a substance burned completely in oxygen
of neutralisation 中和焓变	one mole of water formed from acid and alkali
bond energy 键能	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}).

3 Recap

Topic 5 – key takeaways

1 Enthalpy

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

2 Bond energy

ΔH = bonds broken − bonds made

3 Calorimetry

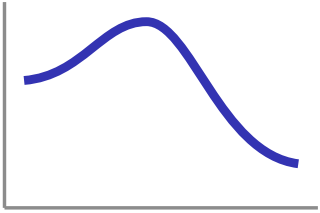
$q = mc\Delta T$; $\Delta H = -q/n$

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