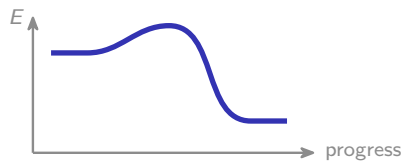
Thermochemistry

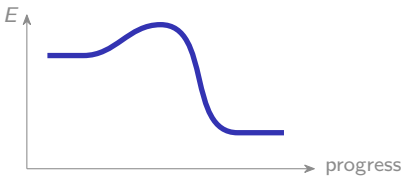
AP Chemistry ▪ Unit 6

AP Chemistry



What we will cover

1. Endothermic and exothermic
2. Energy diagrams and ΔH
3. Calorimetry
4. Phase changes
5. Bond enthalpies
6. Enthalpy of formation
7. Hess's law



1 Endo/exo

Endothermic and exothermic

Energy flows between the **system** 系统 and its **surroundings** 环境:

exothermic 放热: releases energy (warms)

endothermic 吸热: absorbs energy (cools)

Endothermic packs absorb heat from surroundings

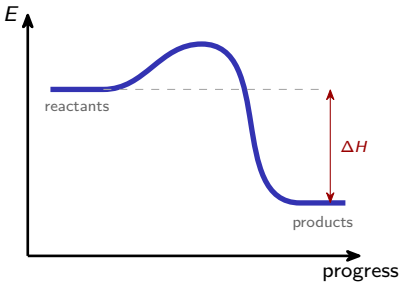
2 Energy diagrams

Energy diagrams and enthalpy

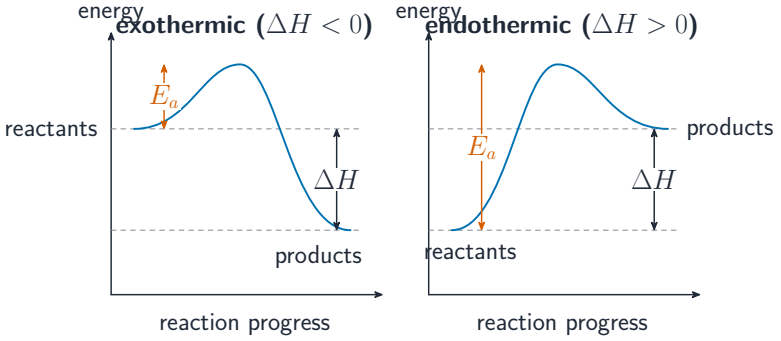
The **enthalpy change** 焓变 ΔH is the heat at constant pressure:

- products lower $\Rightarrow \Delta H < 0$ (exothermic)
- products higher $\Rightarrow \Delta H > 0$ (endothermic)

> ΔH is **extensive** 广延量 – double the moles, double ΔH.



Energy Diagrams



Energy diagrams: exothermic products sit below the reactants, endothermic above

3 Calorimetry

Calorimetry

Specific heat 比热容 c warms $1\,\mathrm{g}$ by 1°:

Heat transferred

$$q = mc\Delta T$$

Calorimetry 量热法: an isolated system, so $q_{rxn} = -q_{water}$.

Worked example

$100\,\mathrm{g}$ water warms 5.0°:

$$q = 100(4.18)(5.0) = 2090\,\mathrm{J}$$

The reaction released $-2090\,\mathrm{J}$ – exothermic. For $0.010\,\mathrm{mol}$: $\Delta H = -209\,\mathrm{kJ/mol}$.

4 Phase changes

Energy of phase changes

A **phase change** 相变 overcomes **intermolecular forces** 分子间作用力, not bonds.

> On a **heating curve** 加热曲线, temperature stays constant during a phase change – energy goes into breaking IMFs.

Heat of vaporization 汽化热 > **heat of fusion** 熔化热 (boiling separates molecules fully).

Phase changes absorb or release latent heat

5 Bond enthalpies

Bond enthalpies

Estimate ΔH from the bonds that change:

Broken minus formed

$$\Delta H \approx \sum(\text{broken}) - \sum(\text{formed})$$

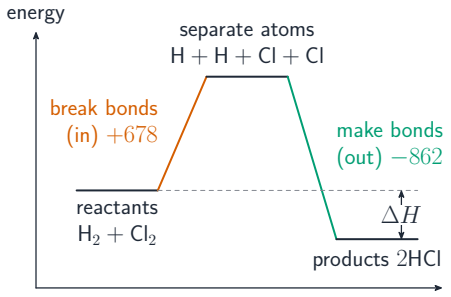
Breaking is endothermic (in), forming exothermic (out).

$H_2 + Cl_2 \rightarrow 2\,HCl$

break H–H (436) + Cl–Cl (243); form 2 H–Cl $(431$ each$)$:

$$(679) - (862) = -183\,\text{kJ}$$

Breaking bonds takes in energy; making bonds



net: $\Delta H = 678 - 862 = -184$ kJ/mol (exothermic)

Breaking bonds takes in energy; making bonds releases it

6 Formation & Hess

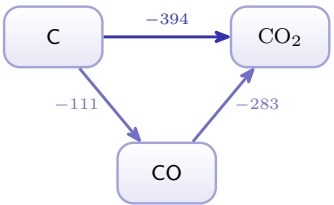
Enthalpy of formation and Hess's law

From a table of ΔH_f°

$$\Delta H^\circ_{rxn} = \sum \Delta H^\circ_f(\text{prod}) - \sum \Delta H^\circ_f(\text{react})$$

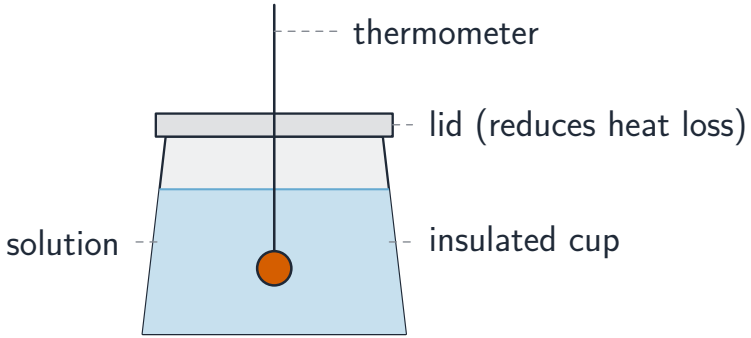
An element in its standard state has $\Delta H_f^\circ = 0$.

Hess's law 赫斯定律: ΔH is a **state function** 状态函数 – add steps. Reverse flips the sign; scale multiplies it.



Same start and end ⇒ same ΔH.

Heat Capacity and Calorimetry



Calorimetry: measure the temperature change of a known mass of solution

Heat, equilibrium, and the enthalpies

heat 热量	flows from hot to cold until objects reach **thermal equilibrium 热平衡**
equilibrium means	equal temperature – not equal energy, and not no motion
enthalpy of reaction 反应焓	ΔH_{rxn} – the heat released or absorbed at constant pressure
the sign	negative for exothermic, positive for endothermic
standard enthalpy of formation 标准生成焓	ΔH_f° – one mole of a compound from its elements in their standard states
why it is useful	$\Delta H_{\text{rxn}} = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$

An element in its standard state has $\Delta H_f^\circ = 0$ by definition. Forgetting that is what breaks the subtraction.

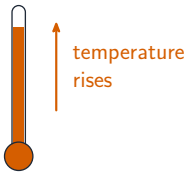
Enthalpy of Formation

formation ΔH_f: elements → 1 mol of the compound

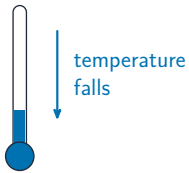
combustion ΔH_c: 1 mol substance + O_2 → $CO_2 + H_2O$

Formation makes a compound from its elements; combustion burns it in oxygen

Endothermic and Exothermic Processes



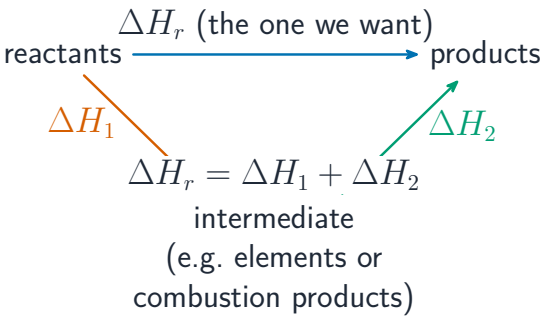
exothermic
heat given out, so the surroundings get hotter
($\Delta H < 0$)



endothermic
heat taken in, so the surroundings get colder
($\Delta H > 0$)

An exothermic reaction warms the surroundings; an endothermic one cools them

Hess's Law



Hess's law: the direct and indirect routes give the same total enthalpy change

Exam tips

- An **exothermic** reaction has $\Delta H < 0$ (feels hot); endothermic has $\Delta H > 0$ – always give ΔH a sign and units.
- In calorimetry use $q = mc\Delta T$ with the mass of the **water/solution**; the reaction releases what the water gains.
- Bond enthalpies: $\Delta H \approx \sum(\text{bonds broken}) - \sum(\text{bonds made})$ – breaking is endothermic, making is exothermic (the classic sign trap).
- **Hess's law**: the total ΔH is the sum of the steps' – reverse a step and flip the sign, scale a step and scale ΔH.
- During a **phase change** the temperature stays constant while energy goes into the forces between particles.

7 Recap

Unit 6 – key takeaways

1 Enthalpy

$\Delta H < 0$ exothermic; break absorbs, form releases

2 Calorimetry

$q = mc\Delta T$; $q_{rxn} = -q_{water}$

3 Phase change

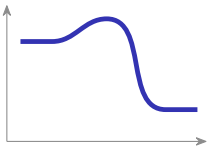
constant T; vaporization $>$ fusion

4 Enthalpy paths

bond enthalpies; ΔH_f°; Hess adds steps

Check yourself

1. If products sit lower than reactants, what is the sign of ΔH?
2. Does breaking a bond absorb or release energy?
3. What happens to temperature during a phase change?
4. Reverse a reaction – what happens to its ΔH?



Answers **1.** negative **2.** absorb **3.** it stays constant **4.** its sign flips