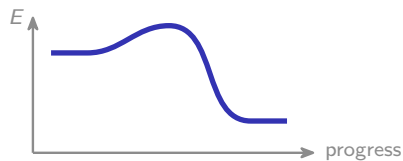
Thermochemistry

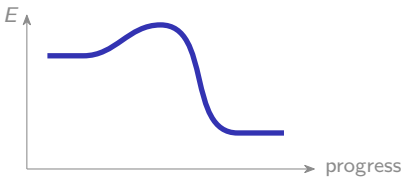
AP Chemistry ▪ Unit 6

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



Endothermic and exothermic

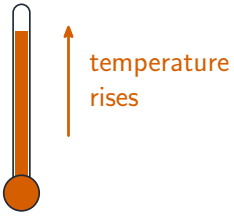
Energy flows between the **system** 系统 and its **surroundings** 环境:

exothermic 放热: releases energy (warms)

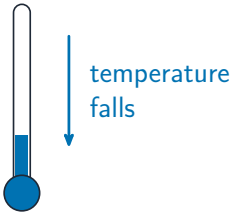
endothermic 吸热: absorbs energy (cools)

Endothermic packs absorb heat from surroundings

Endothermic and exothermic



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



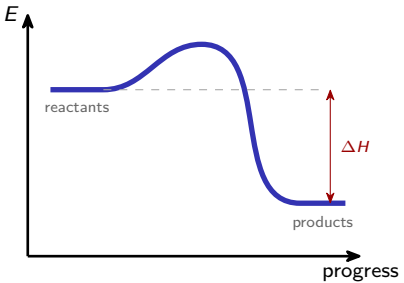
endothermic
heat taken in, so
the surroundings
get colder
($\Delta H > 0$)

Energy diagrams and enthalpy

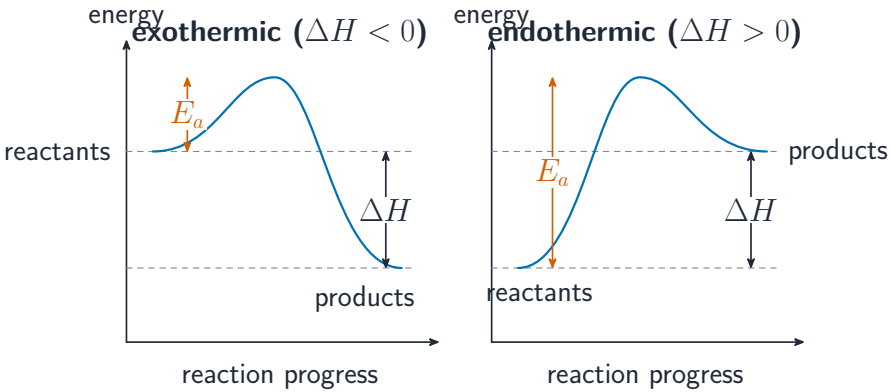
The **enthalpy change** 焓变 ΔH is the heat at constant pressure:

- 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 and enthalpy



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 $q_{rxn} = -q_{water}$.

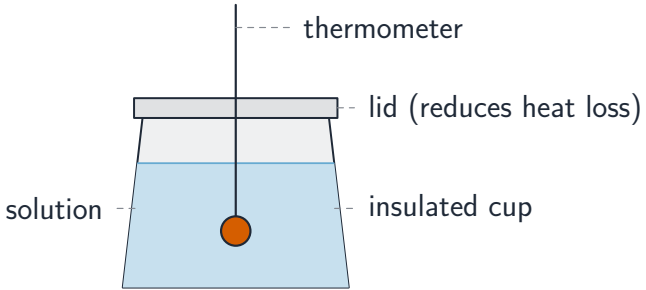
Worked example

$100\,\mathrm{g}$ water warms 5.0°:

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

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

Calorimetry



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

Bond enthalpies

Estimate ΔH from the bonds that change:

Broken minus formed

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

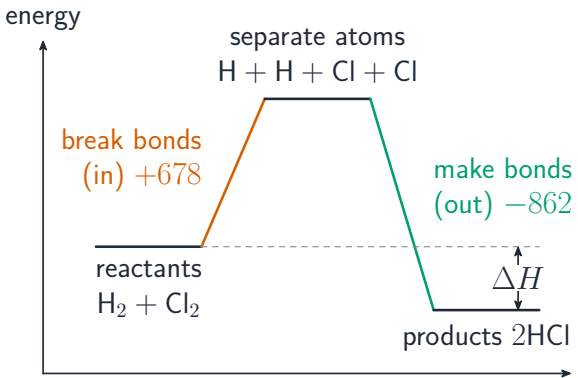
Breaking is endothermic (in), forming exothermic (out).

$H_2 + Cl_2 \rightarrow 2\,HCl$

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

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

Bond enthalpies



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

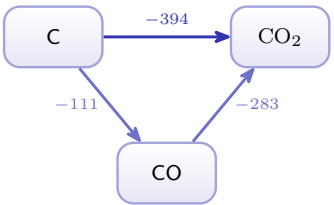
Enthalpy of formation and Hess's law

From a table of ΔH_f°

$$\Delta H_{rxn}^\circ = \sum \Delta H_f^\circ(\text{prod}) - \sum \Delta H_f^\circ(\text{react})$$

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.

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

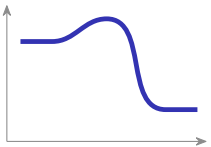
constant T; vaporization $>$ fusion

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