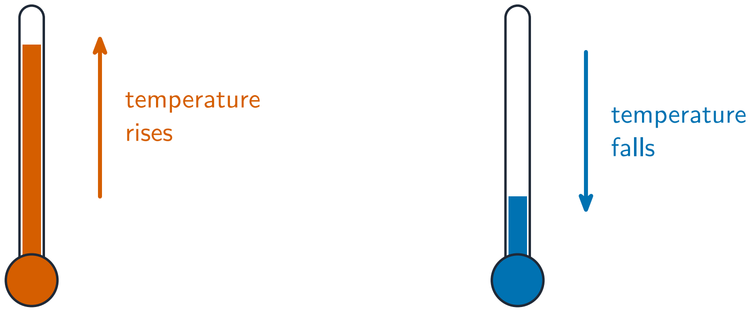
Thermochemistry

AP Chemistry

Endothermic and Exothermic Processes

Thermochemistry 热化学 tracks energy in reactions. A process is **exothermic** 放热 if it releases energy to the surroundings (feels hot, $\Delta H < 0$) and **endothermic** 吸热 if it absorbs energy (feels cold, $\Delta H > 0$). Breaking bonds costs energy; forming bonds releases it –the net decides the sign.



exothermic

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

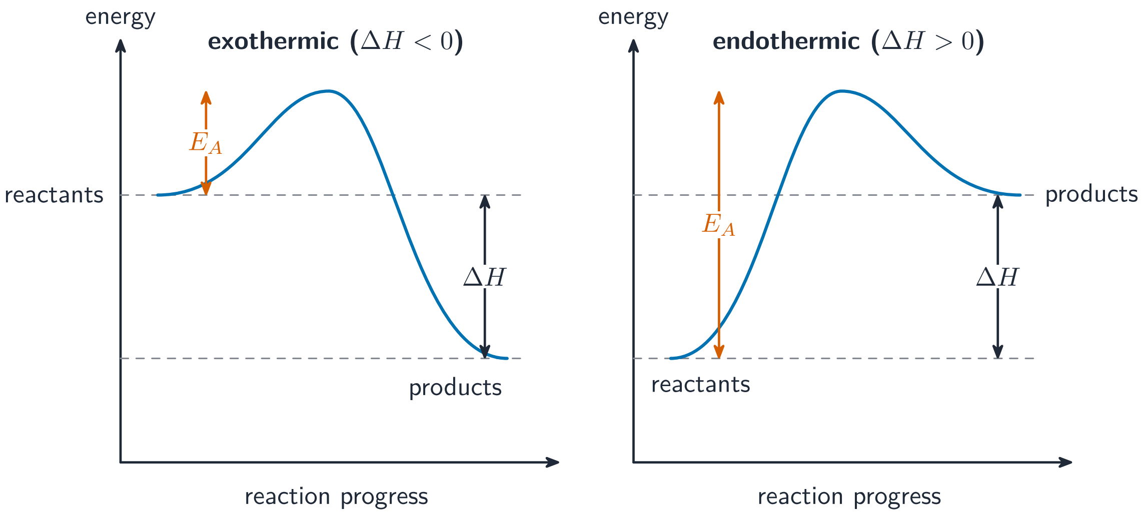
endothermic

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

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

Energy Diagrams

An energy diagram plots energy from reactants to products. Reactants above products means exothermic; below means endothermic. The vertical gap between them is the enthalpy change ΔH.



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

Heat Transfer and Thermal Equilibrium

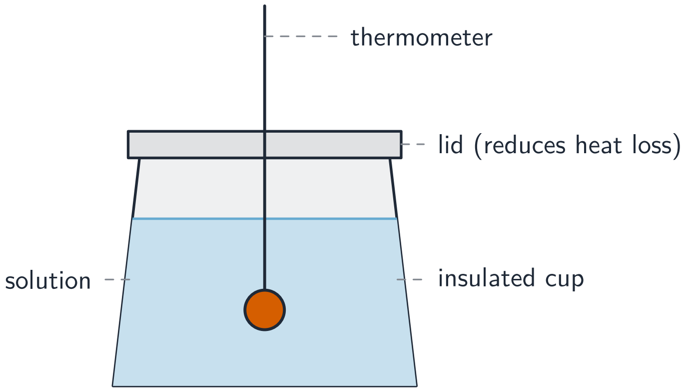
Heat 热量 flows from hot to cold until objects reach **thermal equilibrium** 热平衡 (equal temperature). Energy is conserved: the heat lost by the hot object equals the heat gained by the cold one.

Heat Capacity and Calorimetry

The heat to change a substance's temperature is

$$q = mc\,\Delta T,$$

where c is the **specific heat** 比热容 (energy per gram per degree). **Calorimetry** 量热法 measures a reaction's heat by tracking the temperature change of surrounding water: the heat the water gains equals the heat the reaction releases.



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

Worked example. A reaction in a coffee-cup calorimeter warms 100 g of water by 8.0 °C ($c = 4.18$ J/(g °C)). The heat absorbed by the water is

$$q = mc\,\Delta T = 100 \times 4.18 \times 8.0 = 3.3 \times 10^3 \text{ J}.$$

By energy conservation the reaction **released** this 3.3 kJ, so it is exothermic ($q_{\text{rxn}} = -3.3$ kJ).

Energy of Phase Changes

During a **phase change** 相变 (melting, boiling) the temperature stays **constant** while heat goes into breaking intermolecular forces, not raising kinetic energy. The energy needed is $q = n\,\Delta H_{\text{fus}}$ (melting) or $q = n\,\Delta H_{\text{vap}}$ (boiling) –the flat steps on a heating curve.

Introduction to Enthalpy of Reaction

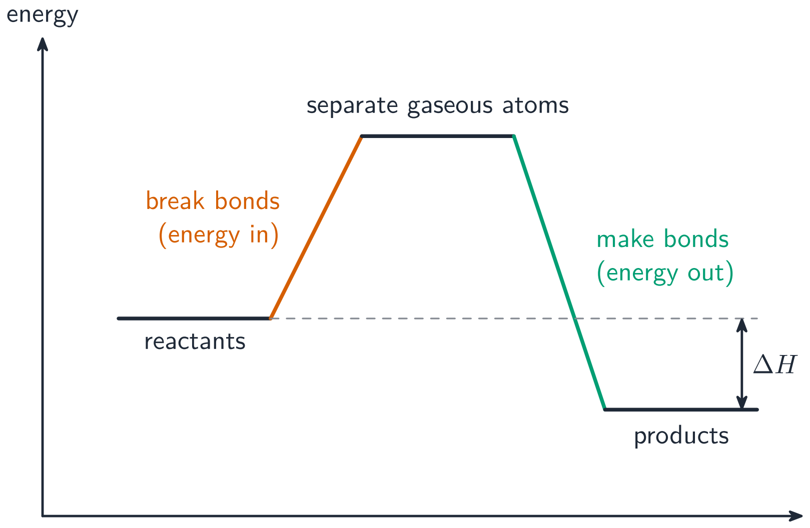
The **enthalpy of reaction** 反应焓 ΔH_{rxn} is the heat released or absorbed at constant pressure. Because enthalpy is a **state function** 状态函数, ΔH depends only on the initial and final states, not the path taken –the key that makes the next three methods work.

Bond Enthalpies

One way to estimate ΔH: sum the energy to **break** all reactant bonds, then subtract the energy **released** forming product bonds:

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

This is an approximation, since bond enthalpies are averages.



Breaking bonds takes in energy; making bonds releases it

Worked example. Estimate ΔH for $H_2 + Cl_2 \rightarrow 2HCl$ using bond enthalpies H–H = 436, Cl–Cl = 242, H–Cl = 431 kJ/mol. Break both reactant bonds ($436 + 242 = 678$) and form two H–Cl bonds ($2 \times 431 = 862$):

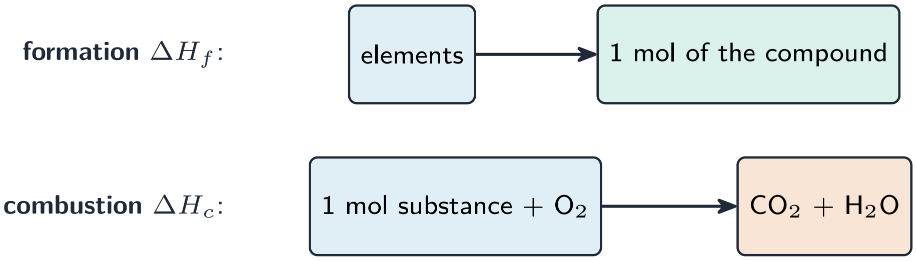
$$\Delta H \approx 678 - 862 = -184 \text{ kJ},$$

exothermic, because the strong H–Cl bonds formed release more than the reactant bonds cost.

Enthalpy of Formation

The **standard enthalpy of formation** 生成焓 ΔH_f° is the enthalpy to make one mole of a compound from its elements (zero for an element in its standard state). Then

$$\Delta H^\circ_{\text{rxn}} = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants}).$$



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

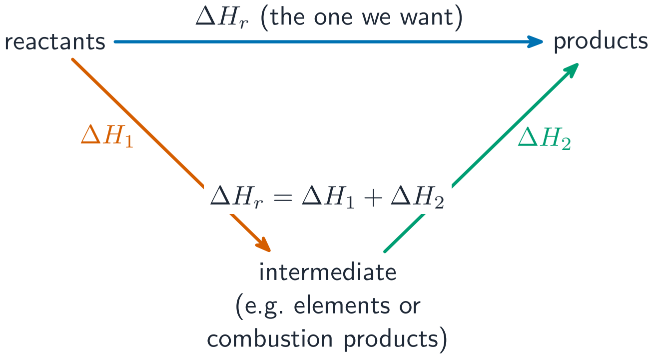
Worked example. Find $\Delta H^\circ_{\text{rxn}}$ for burning methane, $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, given ΔH_f°: $CH_4 = -75$, $CO_2 = -394$, $H_2O = -286$ kJ/mol ($O_2 = 0$). Products minus reactants:

$$\Delta H^\circ_{\text{rxn}} = [-394 + 2(-286)] - [-75 + 0] = -966 + 75 = -891 \text{ kJ},$$

a large release, as expected for a combustion.

Hess's Law

Hess's law 盖斯定律: if a reaction is the sum of several steps, its ΔH is the sum of the steps' ΔH values. Reverse a step and flip the sign; scale a step and scale its ΔH. This lets you find a hard-to-measure ΔH by combining known reactions –a frequent exam calculation.



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.