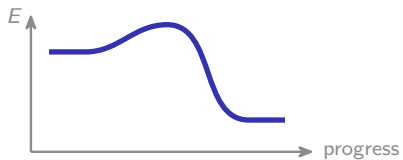


Thermochemistry

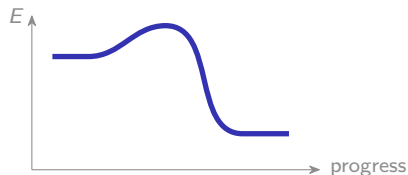
AP Chemistry ■ Unit 6

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**
and its **surroundings** :

exothermic : releases energy (warms)

endothermic : absorbs energy (cools)



*Endothermic packs absorb heat from
surroundings*

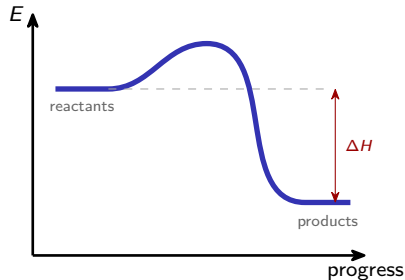


Energy diagrams and enthalpy

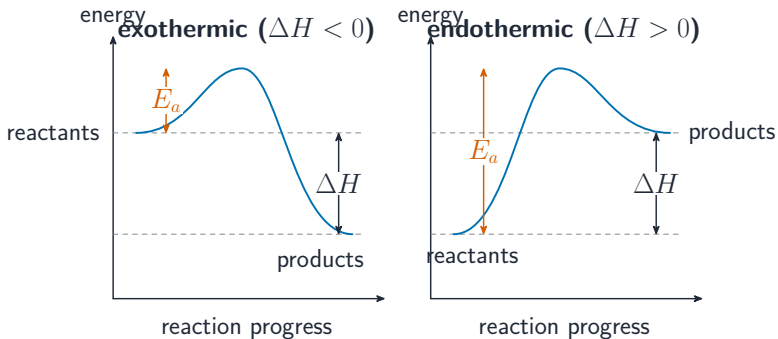
The **enthalpy change** ΔH is the heat at constant pressure:

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

ΔH is **extensive** – double the moles, double ΔH .



Energy Diagrams



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



Calorimetry

Specific heat c warms 1 g by 1° :

Heat transferred

$$q = mc\Delta T$$

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

Worked example

100 g water warms 5.0° :

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

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



Energy of phase changes

A **phase change** overcomes **intermolecular forces**, not bonds.

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

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



Phase changes absorb or release latent heat



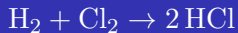
Bond enthalpies

Estimate ΔH from the bonds that change:

Broken minus formed

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

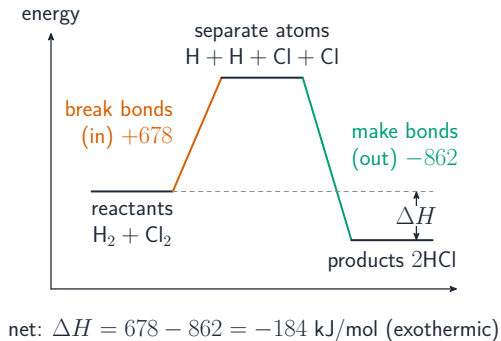
Breaking is endothermic (in), forming exothermic (out).



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



Breaking bonds takes in energy; making bonds releases it



Enthalpy of formation and Hess's law

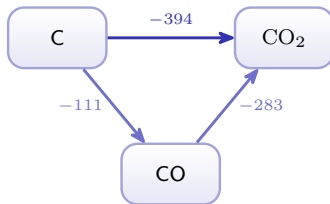
From a table of ΔH_f°

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

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

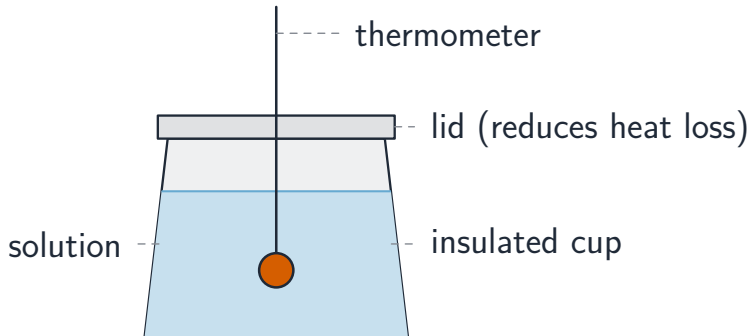
Hess's law : ΔH is a **state function**

– add steps. **Reverse** flips the sign;
scale multiplies it.



Same start
and end $\Rightarrow$ same ΔH .

Heat Capacity and Calorimetry



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

Heat, equilibrium, and the enthalpies

heat

flows from **hot to cold** until objects reach **thermal equilibrium**

equilibrium means

equal temperature – not equal energy, and not no motion

enthalpy of reaction

ΔH_{rxn} – the heat released or absorbed at constant pressure

the sign

negative for exothermic, positive for endothermic

standard enthalpy of formation

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 :

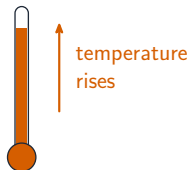
1 mol substance + O₂



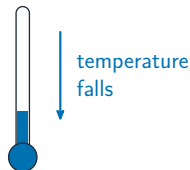
CO₂ + H₂O

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

Endothermic and Exothermic Processes



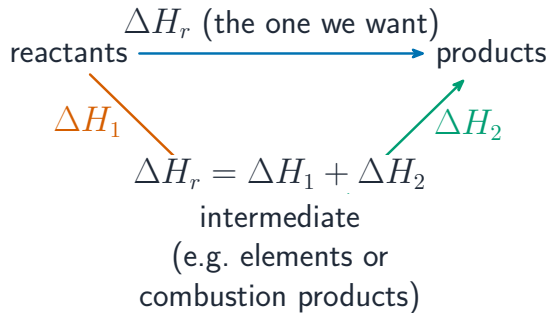
exothermic
heat given out, so
the surroundings
get hotter
($\Delta H < 0$)



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.



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

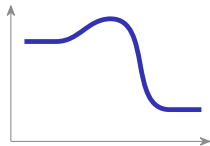
4

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