Energetics

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Energetics

Every chemical reaction involves an energy change. Whether energy is released to the surroundings or absorbed from them depends on the balance between the energy needed to break bonds in the reactants and the energy released when new bonds form in the products.

Exothermic and Endothermic Reactions

An **exothermic reaction** releases heat energy to the surroundings. The temperature of the surroundings rises. The products are at a lower energy than the reactants, so the system has released energy.

An **endothermic reaction** absorbs heat energy from the surroundings. The temperature of the surroundings falls. The products are at a higher energy than the reactants.

Feature	Exothermic	Endothermic
Energy flow	Released to surroundings	Absorbed from surroundings
Temperature change	Surroundings get warmer	Surroundings get cooler
Ħ sign	Negative (Ħ < 0)	Positive (Ħ > 0)
Energy of products vs reactants	Products have less energy	Products have more energy

Examples of exothermic reactions: combustion, neutralisation (acid + alkali), respiration, the reaction of sodium hydroxide pellets with water, freezing and condensation (physical).

Examples of endothermic reactions: thermal decomposition (e.g. CaCOƒCaO + CO), photosynthesis, dissolving potassium nitrate in water, melting and evaporation (physical).

Ħ Notation

The **enthalpy change** (Ħ, pronounced "delta H") of a reaction is the heat energy transferred at constant pressure. It is written with the balanced equation:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890\ \mathrm{kJ\ mol^{-1}}$$

A negative Ħ means heat is released (exothermic). A positive Ħ means heat is absorbed (endothermic).

Bond Energies: Breaking and Forming

All chemical reactions involve:

- 1. Breaking bonds in the reactants — this always **requires** energy and is endothermic.
- 2. Forming bonds in the products — this always **releases** energy and is exothermic.

The overall energy change is the difference between these two:

$$\Delta H = \text{energy to break bonds} - \text{energy released forming bonds}$$

- If more energy is released in bond formation than is absorbed in bond breaking: Ħ is negative ’**exothermic**.
- If more energy is absorbed in bond breaking than is released in bond formation: Ħ is positive ’**endothermic**.

> 🧠 **Remember**
> Breaking bonds costs energy. Forming bonds releases energy. Exothermic reactions form stronger bonds than they break; endothermic reactions break stronger bonds than they form.

Energy Profile Diagrams

An energy profile diagram (reaction coordinate diagram) shows how the energy of the system changes as reactants are converted into products. The x-axis represents the progress of the reaction and the y-axis represents the potential energy of the system.

Exothermic Profile

<EnergyProfileDiagram />

Key features:

- Reactants start at a higher energy level than products.
- The curve rises to a **peak** — the transition state (activated complex) — before falling.
- **Activation energy (E_a):** the difference between the reactant energy level and the peak. This is the energy barrier that must be overcome.
- **Ħ:** the difference in energy between reactants and products. Negative for exothermic reactions (products lower than reactants).

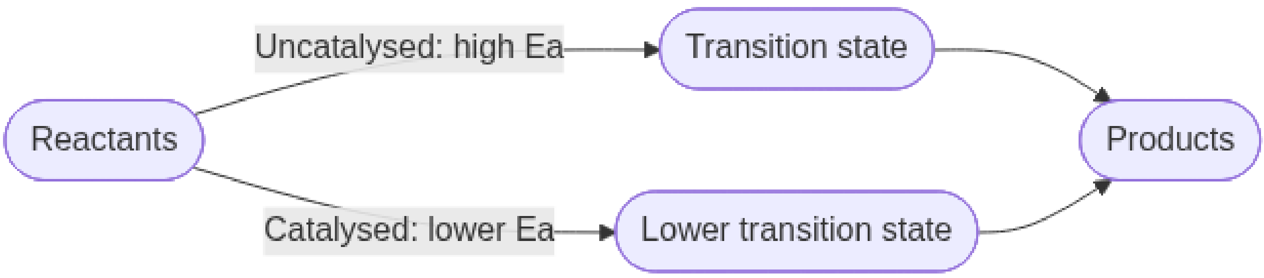
Endothermic Profile

Key features:

- Products end at a higher energy level than reactants.
- The curve still rises over an activation energy barrier before reaching the product level.
- **Ħ** is positive (products higher than reactants).

Effect of a Catalyst

A catalyst provides an alternative pathway with a **lower activation energy** — the peak of the energy profile is lower. The reactant and product energy levels are unchanged, so Ħ is the same. Only the height of the barrier changes.



Catalysed vs uncatalysed reaction

> 📝 **Exam Tip**
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> A catalyst lowers activation energy — it does NOT change Ħ (the overall energy difference between reactants and products). In a diagram question, both the catalysed and uncatalysed curves must reach the same product energy level.

Calculating Energy Changes

The energy transferred in a reaction can be calculated from temperature changes measured in solution using simple calorimetry:

$$q = mc\Delta T$$

where:

- q = heat energy transferred (J)
- m = mass of solution in grams (for dilute aqueous solutions, assume density = 1 g cm^{-3}, so volume in cm^3 = mass in g)
- c = specific heat capacity of water = 4.18 J g^{-1} °C^{-1}
- ΔT = temperature change in °C

For an **exothermic**reaction, the temperature rises (ΔT is positive) and the reaction releases energy, so q is quoted as negative (by convention, energy is released to the surroundings, which is from the reaction).

For an **endothermic**reaction, the temperature falls (ΔT is positive but the reaction absorbs energy).

Standard assumptions in CSEC calculations:

- The density of the solution is 1 g cm^{-3}.
- The specific heat capacity is that of pure water (4.18 J g^{-1} °C^{-1}).
- Heat exchange with the surroundings and the calorimeter is negligible.

Heat of Neutralisation

Heat of neutralisation is the enthalpy change when an acid and a base react to produce one mole of water. For strong acid–strong alkali reactions, it is approximately −57 kJ mol^{-1} regardless of which strong acid and strong alkali are used, because the net reaction is always:

$$\mathrm{H^+(aq) + OH^-(aq) \rightarrow H_2O(l)} \quad \Delta H = -57\ \mathrm{kJ\ mol^{-1}}$$

Heat of Solution

Heat of solution is the enthalpy change when one mole of solute dissolves in excess solvent. Dissolving potassium nitrate in water is endothermic (the solution cools); dissolving sodium hydroxide pellets is exothermic (the solution warms significantly).

📋 **Example**

50 cm^3 of 1.0 mol dm^{-3} HCl was mixed with 50 cm^3 of 1.0 mol dm^{-3} NaOH. The temperature rose from 20.0 °C to 26.8 °C. Calculate the heat released and the molar enthalpy of neutralisation.

Total volume = 100 cm^3; mass of solution = 100 g

$$q = mc\Delta T = 100 \times 4.18 \times (26.8 - 20.0) = 100 \times 4.18 \times 6.8 = 2842 \text{ J} = 2.84 \text{ kJ}$$

Moles of water formed = moles of HCl reacted = 1.0 × 0.050 = 0.050 mol

$$\Delta H = -\frac{q}{n} = -\frac{2.84}{0.050} = -56.8 \text{ kJ mol}^{-1}$$

The negative sign confirms the reaction is exothermic.

Thermometric Titration

In a thermometric titration, temperature is recorded as acid is added progressively to alkali. Temperature rises as neutralisation proceeds (exothermic), reaches a maximum at the equivalence point, then falls as excess acid cools the mixture. Plotting temperature against volume of acid added produces two straight lines — the intersection identifies the exact equivalence point, without requiring a colour indicator.