

Heat in and out of reactions

Chemistry grades 9-10 (GCSE · IGCSE · O Level · Class 9-10) · Lesson

Quick answer: An exothermic reaction releases energy to the surroundings and its temperature rises; an endothermic reaction absorbs energy and its temperature falls. On a reaction profile, ΔH is the difference between products and reactants. Bond energies give $\Delta H = \text{bonds broken} - \text{bonds formed}$, in kJ/mol.

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What are exothermic and endothermic reactions?

Every chemical reaction involves a change in energy. In some reactions energy flows out to the surroundings and you feel warmth. In others energy flows in and the surroundings get cold. Chemists measure this energy change with a thermometer or calculate it from the bonds involved.

Definition

An **exothermic** reaction transfers energy to the surroundings, so the temperature rises. An **endothermic** reaction takes in energy from the surroundings, so the temperature falls. The energy change is written **ΔH** in kJ/mol: negative for exothermic, positive for endothermic.

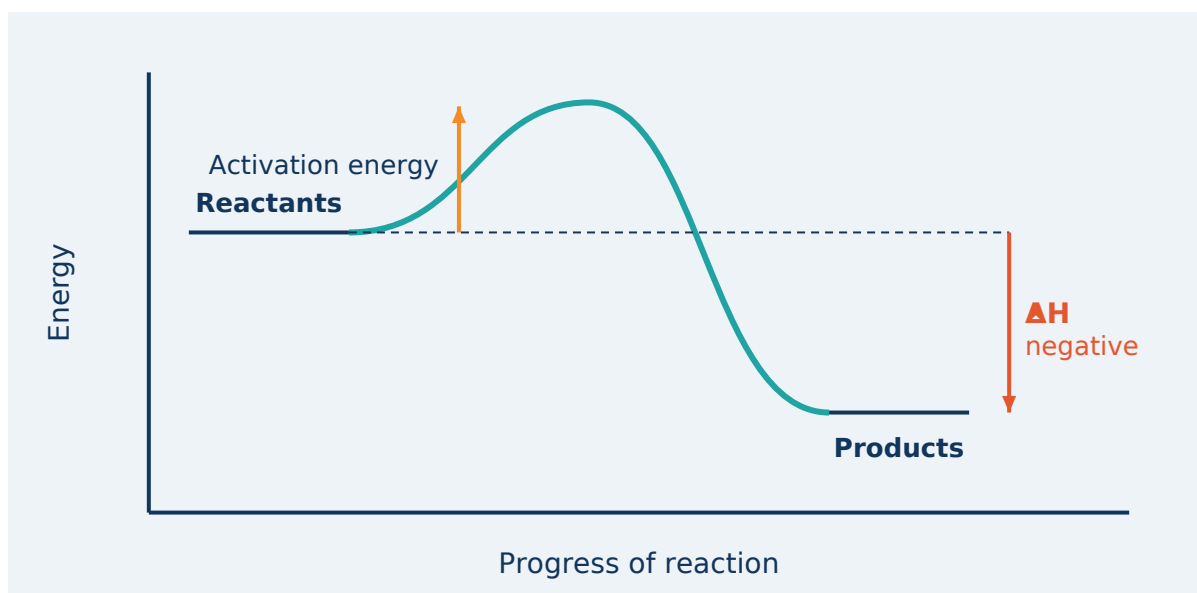
Type	Examples	Everyday use
Exothermic	Burning fuels, neutralisation, respiration, rusting	Hand warmers, self-heating food cans
Endothermic	Thermal decomposition, photosynthesis, dissolving ammonium nitrate	Instant cold packs for sports injuries

Key law

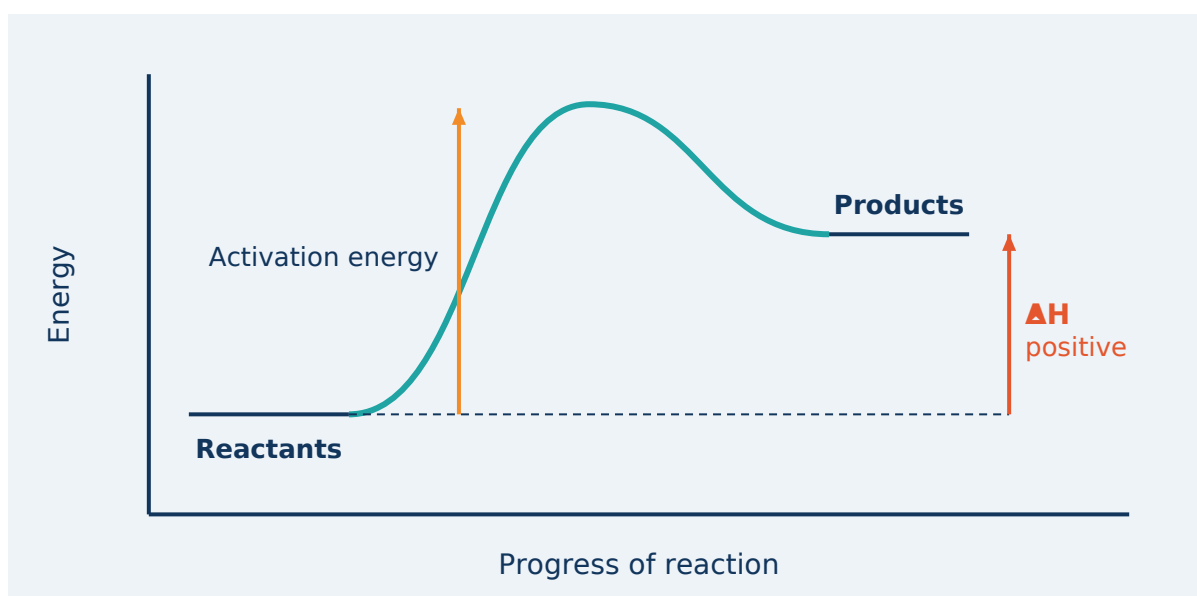
Energy is conserved. $\Delta H = \text{energy of products} - \text{energy of reactants}$. Exothermic: ΔH is negative. Endothermic: ΔH is positive.

What does a reaction profile show?

A reaction profile plots the energy of the particles as a reaction goes from reactants to products. Both kinds of reaction have a hump in the middle. The height of the hump above the reactants is the **activation energy**, the minimum energy colliding particles need so that bonds can break and the reaction can start. This is why a match must be lit before wood burns, even though burning is exothermic.



Exothermic: the products end up lower in energy, so energy is given out.



Endothermic: the products end up higher in energy, so energy is taken in.

Common mistake

Students often think an exothermic reaction needs no energy to start. Every reaction has an activation energy. The overall ΔH is the difference between the products and the reactants, not the height of the hump.

How do bond energies give the energy change?

Atoms are held together by bonds. Breaking a bond always takes in energy, so bond breaking is endothermic. Making a bond always gives out energy, so bond making is exothermic. A reaction is exothermic overall when the energy released making new bonds is greater than the energy needed to break the old ones.

Formula

$$\Delta H = \Sigma(\text{bond energies broken}) - \Sigma(\text{bond energies formed})$$

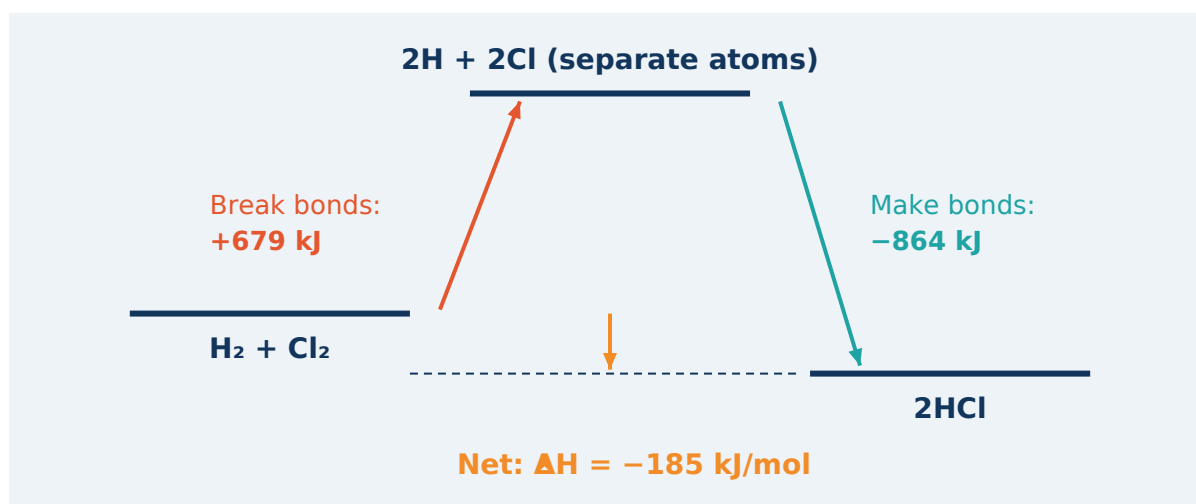
Σ means the sum, and every bond is multiplied by the number of times it appears in the equation. Bond energies are in kJ/mol, so ΔH is in kJ/mol.

Bond	Energy / kJ/mol	Bond	Energy / kJ/mol
H-H	436	O-H	464
Cl-Cl	243	O=O	498
H-Cl	432	C-H	413

Worked example

Calculate ΔH for $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$.

Bonds broken: $1 \times \text{H-H} + 1 \times \text{Cl-Cl} = 436 + 243 = 679$ kJ. Bonds formed: $2 \times \text{H-Cl} = 2 \times 432 = 864$ kJ. $\Delta H = 679 - 864 = -185$ kJ/mol, so the reaction is exothermic.



Breaking bonds costs 679 kJ; making bonds returns 864 kJ; the difference is the energy change.

Worked example

Aarav in Pune burns hydrogen in a rocket demonstration: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$. Calculate ΔH using $\text{H-H} = 436$, $\text{O=O} = 498$ and $\text{O-H} = 464$ kJ/mol.

Broken: $2 \times 436 + 498 = 872 + 498 = 1370$ kJ. Formed: $4 \text{ O-H bonds} = 4 \times 464 = 1856$ kJ. $\Delta H = 1370 - 1856 = -486$ kJ for the equation as written.

How do you measure an energy change with a thermometer?

Water in a cup absorbs the energy from a reaction, so its temperature change tells you how much energy was transferred. The specific heat capacity of water is about $4.2 \text{ J/g}^\circ\text{C}$.

Formula

$$q = m \times c \times \Delta T$$

q = energy transferred (J), m = mass of water (g), $c = 4.2 \text{ J/g}^\circ\text{C}$, ΔT = temperature change ($^\circ\text{C}$). Divide by 1000 to get kJ.

Worked example

Mei adds zinc to 100 g of copper(II) sulfate solution and the temperature rises by 5.0 °C. Calculate the energy released.

$$q = 100 \times 4.2 \times 5.0 = 2100 \text{ J} = \mathbf{2.1 \text{ kJ}}.$$

Lina's tip

Exo = exit: energy exits the reaction and the cup feels hot. Endo = enter: energy enters the reaction and the cup feels cold. Bonds work like Velcro: pulling apart costs energy, sticking together gives it back.

Key points to remember

- Exothermic: temperature rises, ΔH negative; endothermic: temperature falls, ΔH positive
- Activation energy is the minimum energy needed to start a reaction
- Breaking bonds is endothermic; making bonds is exothermic
- $\Delta H = \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$, in kJ/mol
- $q = m \times c \times \Delta T$ with $c = 4.2 \text{ J/g/}^\circ\text{C}$ for water

Heat in and out of reactions: frequently asked questions

How can you tell whether a reaction is exothermic or endothermic?

Measure the temperature of the mixture. If it rises, the reaction is exothermic; if it falls, the reaction is endothermic. A reaction profile also shows it: products lower than reactants means exothermic.

Why does an exothermic reaction such as burning need a spark to start?

The particles must collide with at least the activation energy before bonds can break. The spark supplies that energy, and then the energy released keeps the reaction going.

Why is bond breaking endothermic?

Bonds hold atoms together, so energy must be supplied to pull the atoms apart. Making a bond works in reverse and gives that energy back.

What is the sign of ΔH for an exothermic reaction?

ΔH is negative, because the products have less energy than the reactants. The difference is released to the surroundings.

How do you calculate ΔH from bond energies?

Add the energies of all bonds broken, add the energies of all bonds formed, then subtract: broken minus formed. Remember to multiply by the coefficients in the equation.