Energy Changes in Reactions

Exothermic and endothermic reactions, the enthalpy change and activation energy, reaction pathway diagrams, and calculating the enthalpy change from bond energies.

5.1 Exothermic and endothermic reactions

An exothermic reaction transfers thermal energy to the surroundings, so the temperature of the surroundings increases. An endothermic reaction takes in thermal energy from the surroundings, so the temperature of the surroundings decreases.

	Exothermic	Endothermic
temperature of the surroundings	rises	falls
examples	combustion (burning fuels), neutralisation, respiration, rusting, a metal reacting with an acid	thermal decomposition (such as heating calcium carbonate), photosynthesis, dissolving ammonium nitrate in water

Enthalpy change

The transfer of thermal energy during a reaction is called the **enthalpy change**, ΔH, of the reaction. It is measured in $\mathrm{kJ/mol}$.

- For an **exothermic** reaction, ΔH is **negative**, because the reacting chemicals lose energy to the surroundings.
- For an **endothermic** reaction, ΔH is **positive**, because the reacting chemicals gain energy from the surroundings.

Activation energy

The activation energy, E_a, is the minimum energy that colliding particles must have in order to react.

Even exothermic reactions need some energy to start, which is why a fuel does not burn until it is lit.

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Reaction pathway diagrams

A **reaction pathway diagram** (an energy level diagram) shows the energy of the chemicals as the reaction goes from reactants to products.

To read or draw one:

- The **reactants** are at the left, and the **products** at the right, each shown as a horizontal line at its energy level.
- The **activation energy**, E_a, is the **height from the reactants to the top of the curve**.
- The **enthalpy change**, ΔH, is the **difference in energy between the reactants and the products**, drawn as an arrow between the two levels.
- **Exothermic:** the products are **lower** than the reactants, and the arrow points **down** (ΔH negative).
- **Endothermic:** the products are **higher** than the reactants, and the arrow points **up** (ΔH positive).

NOTE

A **catalyst** gives the reaction a different pathway with a **lower** activation energy. It does **not** change the energy of the reactants or products, so it does **not** change ΔH.

Bond breaking and bond making

- **Breaking bonds is endothermic**: energy must be supplied to pull the atoms apart.
- **Making bonds is exothermic**: energy is released when new bonds form.

The enthalpy change of a reaction = energy needed to break bonds − energy released when bonds are made.

$$\Delta H = (\text{total energy to break the bonds in the reactants}) - (\text{total energy released making the bonds in the products})$$

- If **more energy is released making bonds** than is needed to break bonds, the reaction is **exothermic** (ΔH negative).
- If **more energy is needed to break bonds** than is released, the reaction is **endothermic** (ΔH positive).

TRY IT

Calculate the enthalpy change for $H_2 + Cl_2 \longrightarrow 2\,HCl$. Bond energies in kJ/mol: H−H 436, Cl−Cl 242, H−Cl 431.

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Bonds **broken**: one $\mathrm{H{-}H}$ and one $\mathrm{Cl{-}Cl}$: $436 + 242 = 678$ kJ. Bonds **made**: two $\mathrm{H{-}Cl}$: $2 \times 431 = 862$ kJ.

$$\Delta H = 678 - 862 = -184 \text{ kJ/mol}$$

The sign is negative, so the reaction is **exothermic**.

TRY IT

Calculate the enthalpy change for the combustion of methane: $\mathrm{CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O}$. Bond energies in kJ/mol: C−H 412, O=O 496, C=O 805, O−H 463.

Bonds **broken**: four $\mathrm{C{-}H}$ and two $\mathrm{O{=}O}$: $(4 \times 412) + (2 \times 496) = 1648 + 992 = 2640$ kJ. Bonds **made**: two $\mathrm{C{=}O}$ and four $\mathrm{O{-}H}$: $(2 \times 805) + (4 \times 463) = 1610 + 1852 = 3462$ kJ.

$$\Delta H = 2640 - 3462 = -822 \text{ kJ/mol}$$

NOTE

Before you calculate, **draw out the displayed formulae** so that you can count every bond, including the double bonds. Write "broken" and "made" separately, and give the sign of ΔH at the end.

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