

5.1 Enthalpy change

Name: _____ Class: _____ Date: _____

Total: 25 marks

KEY WORDS enthalpy change • exothermic • bond energy
 • reaction pathway diagram • activation energy

Objective

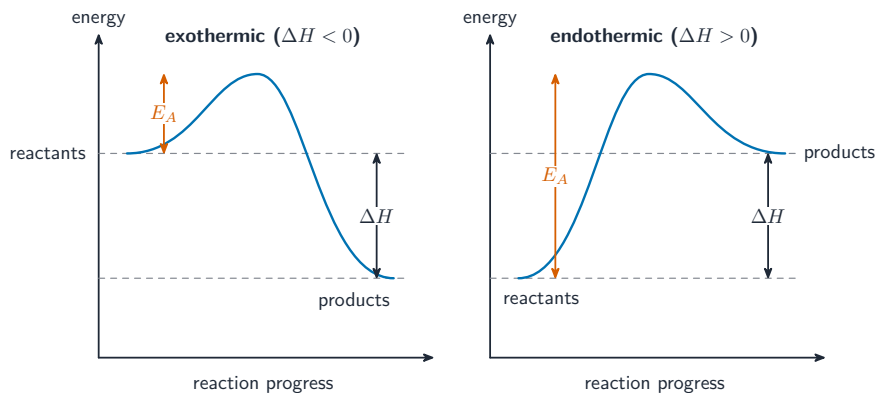
Build the skills to answer exam questions on **enthalpy change** – **exothermic** and **endothermic** reactions, the standard enthalpy changes, **bond energy** calculations, and $q = mc\Delta T$.

You must be able to:

- interpret reaction pathway diagrams and define the standard enthalpy changes
- calculate ΔH_r from bond energies
- calculate enthalpy changes from experiment using $q = mc\Delta T$ and $\Delta H = -mc\Delta T/n$

1 Key knowledge and worked examples

■ Exothermic and endothermic

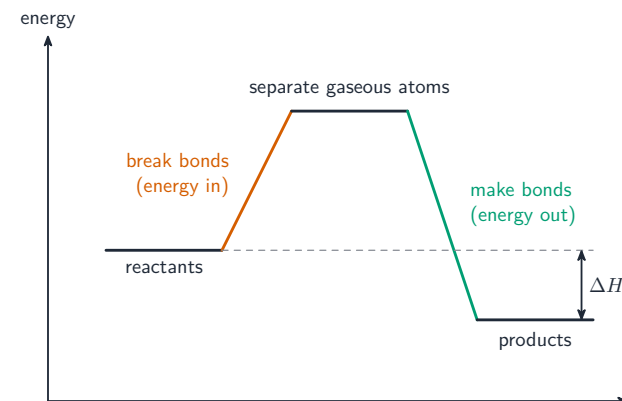


Exothermic: products lower, $\Delta H < 0$. Endothermic: products higher, $\Delta H > 0$. The hill height is the activation energy

Standard conditions: 298 K, 101 kPa ($^\ominus$). Know ΔH_f (formation), ΔH_c (combustion), ΔH_{neut} (neutralisation).

■ Bond energies

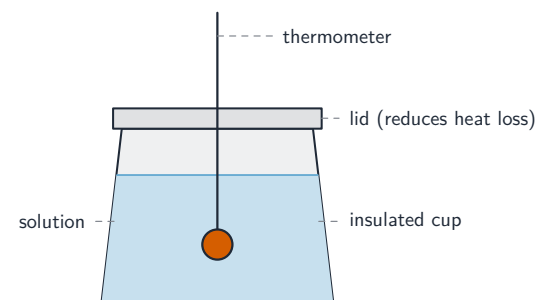
$$\Delta H_r = \sum(\text{bonds broken}) - \sum(\text{bonds made})$$



Breaking bonds takes energy in; making bonds gives energy out

Example. $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$ (H–H 436, Cl–Cl 242, H–Cl 431): $\Delta H = (436 + 242) - (2 \times 431) = -184 \text{ kJ mol}^{-1}$.

■ Measuring ΔH ($q = mc\Delta T$)



Measure the temperature change of a known mass of solution

$q = mc\Delta T$; $\Delta H = -q/n$. **Example.** 0.50 g methanol (M_r 32) heats 100 g water by 18 $^\circ\text{C}$ ($c = 4.18$): $q = 100 \times 4.18 \times 18 = 7520 \text{ J}$; $n = 0.50/32 = 0.0156$; $\Delta H_c = -7520/0.0156 \approx -480 \text{ kJ mol}^{-1}$.

2 Practice

2.1 State whether ΔH is positive or negative for an exothermic reaction, and what this means for the products' energy. [2]

2.2 Define the **enthalpy change of combustion**. [2]

3 Exam-style questions

3.1 In an endothermic reaction: [1]

- ☐ **A** ΔH is negative and heat is given out
B ΔH is positive and heat is taken in
C the products are lower in energy
D no bonds are broken

3.2 Breaking a chemical bond is: [1]

- ☐ **A** exothermic (releases energy)
B endothermic (takes in energy)
C neither
D always $+431 \text{ kJ mol}^{-1}$

3.3 For $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$, use bond energies (kJ mol^{-1} : $\text{N}\equiv\text{N}$ 945, $\text{H}-\text{H}$ 436, $\text{N}-\text{H}$ 391) to calculate ΔH_r . [3]

3.4 Burning 0.32 g of a fuel ($M_r = 32$) raises the temperature of 200 g of water by 9.0°C ($c = 4.18$).

(a) Calculate the heat released, q , in J. [1]

(b) Calculate the enthalpy change of combustion per mole. [3]

3.5 A student burns 0.780 g of propan-1-ol, $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$, and uses the heat released to warm 150 g of water from 21.0°C to 48.5°C . ($c = 4.18 \text{ J g}^{-1}\text{K}^{-1}$; $M_r = 60.0$)

(a) Calculate the enthalpy change of combustion of propan-1-ol from these results. [4]

(b) The data book value is $-2021 \text{ kJ mol}^{-1}$. Suggest **two** reasons for the difference, and state which is likely to be the larger. [3]

(c) Explain why the value is negative, and what that says about the bonds broken and the bonds made. [3]

(d) State **two** reasons why the value obtained this way is not a *standard* enthalpy change of combustion. [2]
