A-LEVEL CHEMISTRY • EXERCISE SHEET

5.1 Enthalpy change

Name: _____ Class: _____ Date: _____ **Total: 25 marks**

KEY WORDS enthalpy change 焓变 • exothermic 放热 • bond energy 键能 • reaction pathway diagram 反应路径图 • activation energy 活化能

Objective

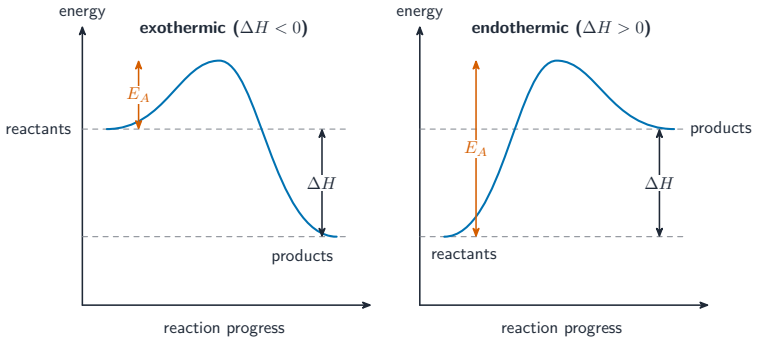
Build the skills to answer exam questions on **enthalpy change** 焓变 — **exothermic** 放热 and **endothermic** 吸热 reactions, the standard enthalpy changes, **bond energy** 键能 calculations, and $q = mc\Delta T$.

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 Worked examples

Exothermic and endothermic

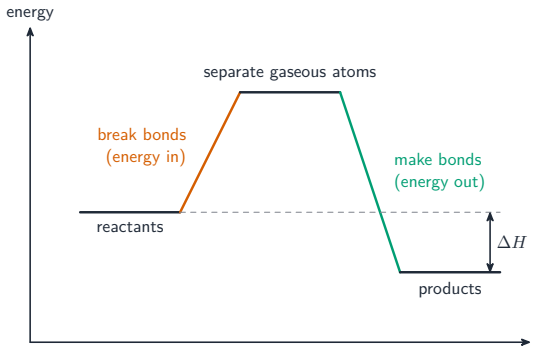


Exothermic: products lower, $\Delta H < 0$. Endothermic: products higher, $\Delta H > 0$. The hill height is the activation energy

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

Bond energies

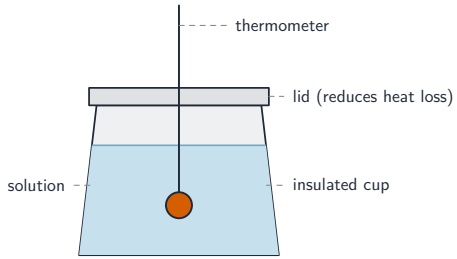
$$\Delta H_r = \sum(\text{bonds broken}) - \sum(\text{bonds made})$$



Breaking bonds takes energy in; making bonds gives energy out

Example. $H_2 + Cl_2 \rightarrow 2HCl$ (H–H 436, Cl–Cl 242, H–Cl 431): $\Delta H = (436 + 242) - (2 \times 431) = -184$ 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 °C ($c = 4.18$): $q = 100 \times 4.18 \times 18 = 7520$ J; $n = 0.50/32 = 0.0156$; $\Delta H_c = -7520/0.0156 \approx -480$ 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 kJ mol^{-1}

3.3 For $N_2 + 3H_2 \rightarrow 2NH_3$, use bond energies (kJ mol^{-1}: N≡N 945, H–H 436, N–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, $CH_3CH_2CH_2OH$, and uses the heat released to warm 150 g of water from 21.0 °C to 48.5 °C. ($c = 4.18$ J g^{-1}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 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]

Solutions

2.1 ΔH is negative; the products have less energy than the reactants.

2.2 the enthalpy change when one mole of a substance is burned completely in oxygen under standard conditions.

3.1 B. Endothermic = ΔH positive, heat taken in.

3.2 B. Breaking a bond requires energy (endothermic).

3.3 broken $= 945 + 3(436) = 2253$; made $= 6 \times 391 = 2346$; $\Delta H = 2253 - 2346 = -93$ kJ mol^{-1}.

3.4 (a) $q = 200 \times 4.18 \times 9.0 = 7524$ J.
(b) $n = 0.32/32 = 0.010$ mol; $\Delta H_c = -q/n = -7524/0.010 = -752$ kJ mol^{-1} (≈ -750).

3.5 (a) $q = mc\Delta T = 150 \times 4.18 \times 27.5 = 17\,240$ J $= 17.2$ kJ; $n = \dfrac{0.780}{60.0} = 0.0130$ mol;
$\Delta H_c = -\dfrac{17.2}{0.0130} = -1.33 \times 10^3$ kJ mol^{-1}.
(b) Heat is lost to the surroundings and to the apparatus instead of warming the water; combustion is incomplete, or some alcohol evaporates without burning; in an open calorimeter the **heat loss** is normally much the larger effect.
(c) Negative means energy is released to the surroundings, i.e. the reaction is exothermic; more energy is released when the bonds in CO_2 and H_2O form than is absorbed breaking the bonds in the alcohol and oxygen; so the products lie at a lower enthalpy than the reactants.
(d) The conditions are not standard — the temperature changes throughout and the pressure is not controlled at 100 kPa; and the alcohol is not necessarily burned completely to CO_2 and H_2O.