

Exercise walk-through

5.1

Enthalpy change, ΔH

A-Level Chemistry

Enthalpy change ΔH

Enthalpy change ΔH is the heat transferred at constant pressure.

exothermic

$\Delta H < 0$ – heat given
out; products more stable

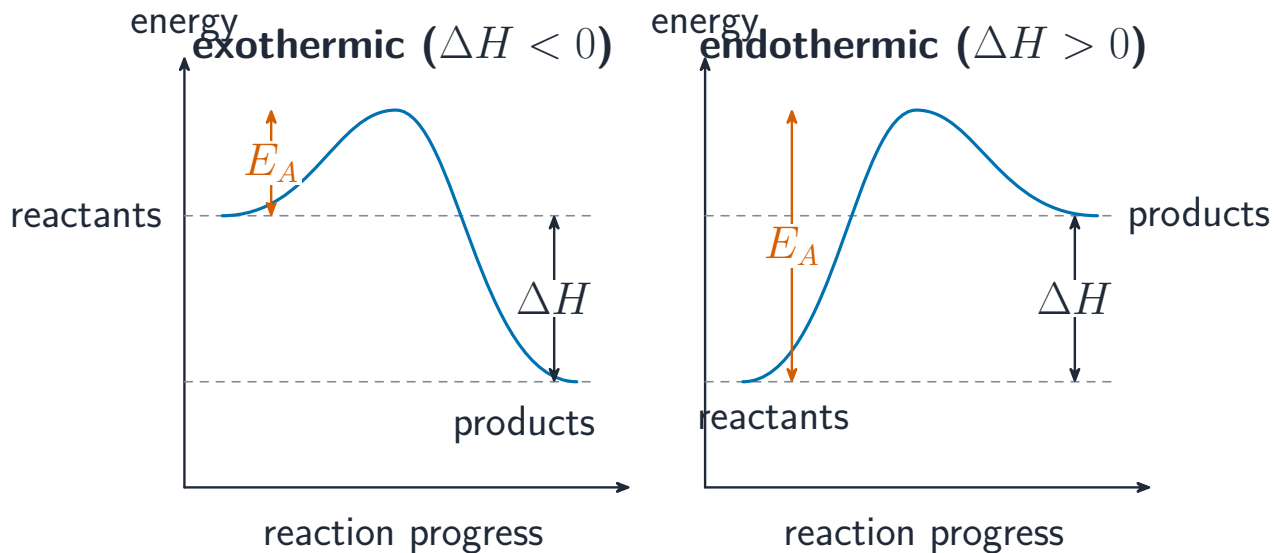
endothermic

$\Delta H > 0$ – heat
taken *in*; a cold pack

Reaction pathway diagrams

- exothermic: products sit **lower** than reactants ($\Delta H < 0$).
- endothermic: products sit **higher** than reactants ($\Delta H > 0$).

Reaction pathway diagrams, in detail



endothermic: products sit **higher** than reactants ($\Delta H > 0$).

Types of enthalpy change

$\Delta H_r^\ominus$ **reaction** – for the
equation as written

$\Delta H_c^\ominus$ **combustion** – 1
mol burnt in oxygen

$\Delta H_f^\ominus$ **formation** – 1
mol of compound from elements

$\Delta H_{\text{neut}}^\ominus$ **neutralisation**
– 1 mol of water formed

Standard conditions

($^\ominus$): 298 K, 101 kPa, each substance in its normal state.

Exercises 2.1·2.2

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]

Exercises 2.1·2.2

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

2.1

- ΔH is negative; the products have less energy than the reactants
-

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

2.2

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

Exercises 3.1·3.2

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 $+431 \text{ kJ mol}^{-1}$

Exercises 3.1·3.2

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[1]

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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

3.1 B

- Endothermic = ΔH positive, heat taken in

3.2 Breaking a chemical bond is:

[1]

☐

A exothermic (releases energy)

B endothermic (takes in energy)

C neither

D always $+431 \text{ kJ mol}^{-1}$

3.2 B

- Breaking a bond requires energy (endothermic)

Exercise 3.3

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]

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3.3

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

Exercise 3.4 (a)

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]

Exercise 3.4 (a)

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]

(a)

- $q = 200 \times 4.18 \times 9.0 = 7524 \text{ J}$

Exercise 3.4 (b)

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

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

[3]

Exercise 3.4 (b)

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

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

(b)

- $n = 0.32/32 = 0.010$ mol; $\Delta H_c = -q/n = -7524/0.010 = -752$ kJ mol⁻¹ (≈ -750)

Exercise 3.5 (a)

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]

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(a)

- $q = mc\Delta T = 150 \times 4.18 \times 27.5 = 17\,240 \text{ J} = 17.2 \text{ kJ}; n = \frac{0.780}{60.0} = 0.0130 \text{ mol};$
$$\Delta H_c = -\frac{17.2}{0.0130} = -1.33 \times 10^3 \text{ kJ mol}^{-1}$$

Exercise 3.5 (b)

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. ($c = 4.18 \text{ J g}^{-1}\text{K}^{-1}$; $M_r = 60.0$)

(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]

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(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

Exercise 3.5 (c)

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

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(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

Exercise 3.5 (d)

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

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(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