Exercise walk-through

Enthalpy change, ΔH ▪ 5.1

eXercise

You must be able to

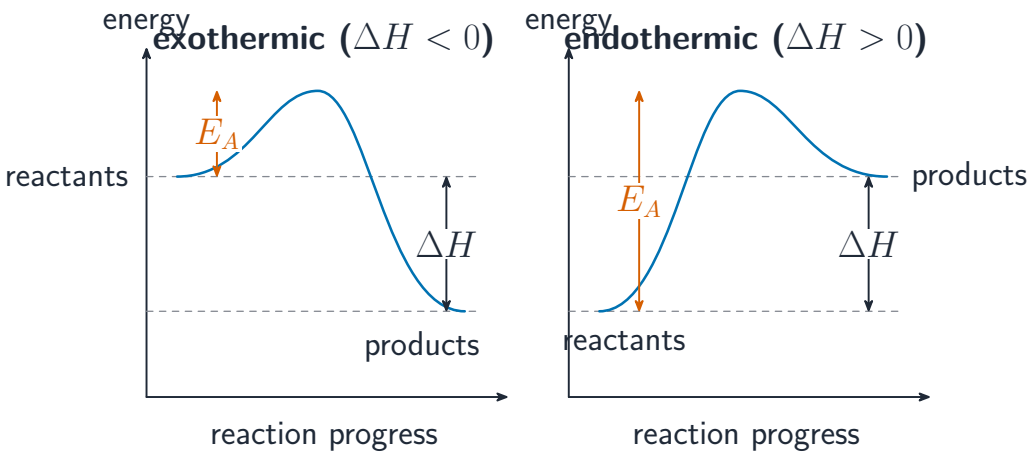
- interpret reaction pathway diagrams and define the standard enthalpy changes
- calculate ΔH_r from bond energies
- calculate enthalpy changes from experiment using $q = mc\Delta T$ and $\Delta H = -mc\Delta T/n$

Method — 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).

Method — Exothermic and endothermic

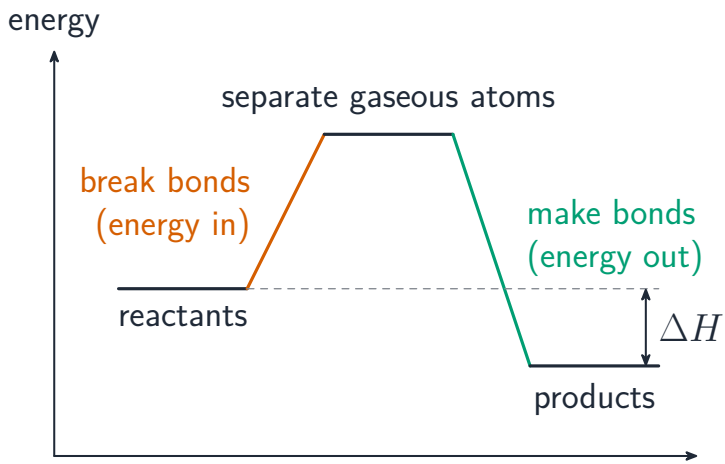


Method — Bond energies

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

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}.

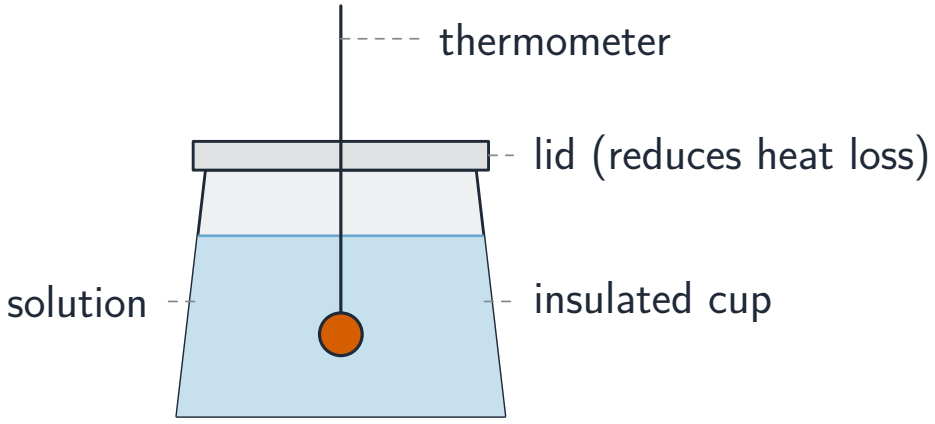
Method — Bond energies



Method — Measuring ΔH ($q = mc\Delta T$)

$q = mc\Delta T$; $\Delta H = -q/n$. **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}.

Method — Measuring ΔH ($q = mc\Delta T$)



An insulated cup with a thermometer measuring the temperature change of a solution

Practice 2.1 ▪ 2 marks

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

Solution 2.1

Method: Exothermic and endothermic

Worked steps

ΔH is negative B1; the products have less energy than the reactants B1.

Practice 2.2 ▪ 2 marks

Define the **enthalpy change of combustion**.

Solution 2.2

Worked steps

the enthalpy change when one mole of a substance B1 is burned completely in oxygen under standard conditions B1.

Exam-style 3.1 ▪ 1 mark

In an endothermic reaction:

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

Solution 3.1

Method: Exothermic and endothermic

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

Worked steps

B. Endothermic = ΔH positive, heat taken in.

Exam-style 3.2 ▪ 1 mark

Breaking a chemical bond is:

A exothermic (releases energy)

B endothermic (takes in energy)

C neither

D always $+431$ kJ mol^{1}

Solution 3.2

Method: Bond energies

A exothermic (releases energy)

B endothermic (takes in energy) ✓

C neither

D always +431 kJ mol ^{1}

Worked steps

B. Breaking a bond requires energy (endothermic).

Exam-style 3.3 ▪ 3 marks

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.

Solution 3.3

Method: Bond energies

Worked steps

broken $= 945 + 3(436) = 2253$; made $= 6 \times 391 = 2346$ M1 M1; $\Delta H = 2253 - 2346 = -93$ kJ mol^{-1} A1.

Exam-style 3.4 ▪ 1 mark

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.

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

Solution 3.4

Method: Measuring ΔH ($q = mc\Delta T$)

Worked steps

(a) $q = 200 \times 4.18 \times 9.0 = 7524$ J **B1**.

(b) $n = 0.32/32 = 0.010$ mol **M1**; $\Delta H_c = -q/n = -7524/0.010$ **M1** $= -752$ kJ mol^{-1} (≈ -750) **A1**.