

Exercise walk-through

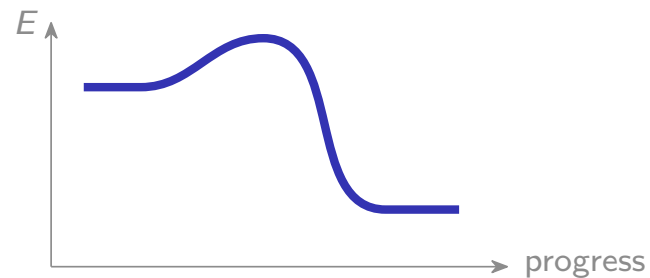
6.6

Introduction to Enthalpy of Reaction

AP Chemistry

What we will cover

- 1 Endothermic and exothermic
- 2 Energy diagrams and ΔH
- 3 Calorimetry
- 4 Phase changes
- 5 Bond enthalpies
- 6 Enthalpy of formation
- 7 Hess's law



Heat, equilibrium, and the enthalpies

heat

flows from **hot to cold** until objects reach **thermal equilibrium**

equilibrium means

equal temperature – not equal energy, and not no motion

enthalpy of reaction

ΔH_{rxn} – the heat released or absorbed at constant pressure

the sign

negative for exothermic, positive for endothermic

standard enthalpy of formation

ΔH_f° – one mole of a compound from its elements in their standard states

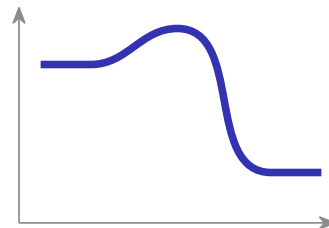
why it is useful

$$\Delta H_{\text{rxn}} = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$$

An element in its standard state has $\Delta H_f^\circ = 0$ **by definition**. Forgetting that is what breaks the subtraction.

Check yourself

- 1 If products sit lower than reactants, what is the sign of ΔH ?
- 2 Does breaking a bond absorb or release energy?
- 3 What happens to temperature during a phase change?
- 4 Reverse a reaction – what happens to its ΔH ?



Answers 1. negative 2. absorb 3. it stays constant 4. its sign flips

Exercises 2.1-2.3

2.1 What does a negative ΔH_{rxn} mean?

[1]

2.2 A reaction releases 150 kJ/mol. How much for 3 mol?

[1]

2.3 How does reversing a reaction change ΔH ?

[1]

Exercises 2.1-2.3

2.1 What does a negative ΔH_{rxn} mean? [1]

2.1

- The reaction is exothermic (releases heat)

2.2 A reaction releases 150 kJ/mol. How much for 3 mol? [1]

2.2

- 450 kJ

2.3 How does reversing a reaction change ΔH ? [1]

2.3

- It flips the sign

Exercise 3.1

3.1 If a reaction as written has $\Delta H = -100$ kJ, then doubling all coefficients gives [1]

☐

A -50 kJ

B -100 kJ

C -200 kJ

D $+100$ kJ

Exercise 3.1

3.1 If a reaction as written has $\Delta H = -100 \text{ kJ}$, then doubling all coefficients gives [1]

☐

A -50 kJ

B -100 kJ

C -200 kJ

D $+100 \text{ kJ}$

3.1 C

- ΔH scales with the coefficients, so it doubles to -200 kJ

Exercise 3.2 (a)

3.2 The reaction $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$ has $\Delta H = -92 \text{ kJ}$.

(a) State whether it is exo- or endothermic.

[1]

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[1]

(a)

- Exothermic

Exercise 3.2 (b)

3.2 The reaction $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$ has $\Delta H = -92 \text{ kJ}$.

(b) Find the heat released when 1.0 mol NH_3 forms.

[2]

Exercise 3.2 (b)

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(b) Find the heat released when 1.0 mol NH_3 forms.

[2]

(b)

- ΔH is -92 kJ per 2 mol NH_3 , so per 1 mol it is -46 kJ

Exercise 3.3

3.3 Combustion of methane releases 890 kJ/mol. Find the heat released by burning 8.0 g of methane ($M = 16$ g/mol). [3]

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3.3

- $8.0 \text{ g} \div 16 = 0.50 \text{ mol}$; heat = $0.50 \times 890 = 445 \text{ kJ}$