

Exercise walk-through

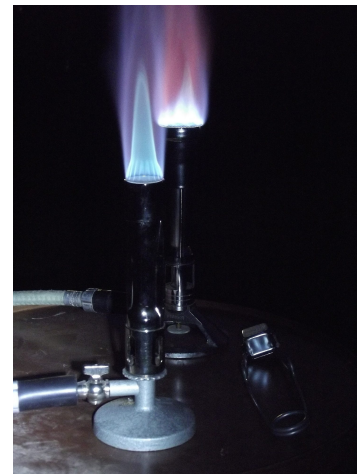
5.2

Hess's law

A-Level Chemistry

What we will cover

- 1 Enthalpy change ΔH
- 2 Types of enthalpy change
- 3 Bond energy
- 4 Measuring ΔH
- 5 Hess's law



burning fuel is exothermic

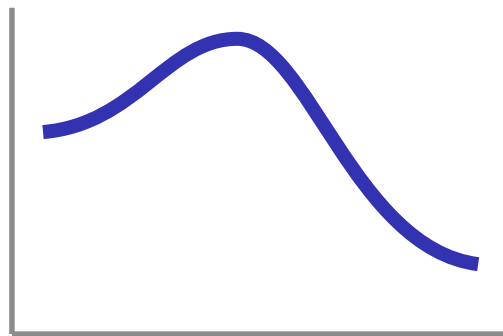
The named enthalpy changes, and Hess's law

a reaction pathway diagram	shows the energy of reactants and products, and the energy hill between them – the activation energy
of formation	one mole of a compound from its elements in their standard states
of combustion	one mole of a substance burned completely in oxygen
of neutralisation	one mole of water formed from acid and alkali
bond energy	the energy to break one mole of a particular bond in the gas state – so it is always positive
Hess's law	the enthalpy change is the same by any route – so an unmeasurable ΔH can be got from measurable ones

Hess's law is used two ways: from enthalpies of **formation** (products minus reactants) and from enthalpies of **combustion** (reactants minus products). The two subtractions run opposite ways.

Check yourself

- 1 Sign of ΔH for an exothermic reaction?
- 2 Does breaking a bond take in or give out energy?
- 3 Write the equation for heat transferred.
- 4 What does Hess's law say about the route?



Answers 1. negative 2. takes in energy 3. $q = mc\Delta T$ 4. it is the same by any route

Exercises 2.1·2.2

2.1 State **Hess's law**.

[1]

2.2 Give **one** reason Hess's law is useful (something it lets you find).

[1]

Exercises 2.1·2.2

2.1 State **Hess's law**.

[1]

2.1

- the total enthalpy change of a reaction is the same regardless of the route taken

2.2 Give **one** reason Hess's law is useful (something it lets you find).

[1]

2.2

- e.g. it lets you find an enthalpy change that cannot be measured directly / calculate ΔH_r from formation or combustion data

Exercises 3.1·3.2

3.1 Hess's law is a consequence of:

[1]

☐

- A the conservation of mass
- B the conservation of energy
- C the ideal gas equation
- D Avogadro's law

3.2 ΔH_r from formation data is calculated as:

[1]

☐

- A $\sum \Delta H_f(\text{reactants}) - \sum \Delta H_f(\text{products})$
- B $\sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$
- C $\sum \Delta H_c(\text{products}) - \sum \Delta H_c(\text{reactants})$
- D $mc\Delta T$

Exercises 3.1·3.2

3.1 Hess's law is a consequence of:

[1]

☐

- A the conservation of mass
- B the conservation of energy
- C the ideal gas equation
- D Avogadro's law

3.1 B

- Hess's law follows from the conservation of energy

3.2 ΔH_r from formation data is calculated as:

[1]

☐

- A $\sum \Delta H_f(\text{reactants}) - \sum \Delta H_f(\text{products})$
- B $\sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$
- C $\sum \Delta H_c(\text{products}) - \sum \Delta H_c(\text{reactants})$
- D $mc\Delta T$

3.2 B

- $\sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$

Exercises 3.3·3.4

3.3 For $\text{C}_2\text{H}_4 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6$, use ΔH_f (kJ mol^{-1} : $\text{C}_2\text{H}_4 +52$, $\text{C}_2\text{H}_6 -85$, $\text{H}_2 0$) to calculate ΔH_r . [3]

3.4 Explain why Hess's law allows the enthalpy change of formation of a compound to be found even when the reaction cannot be carried out directly. [2]

Exercises 3.3·3.4

3.3 For $\text{C}_2\text{H}_4 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6$, use ΔH_f (kJ mol^{-1} : $\text{C}_2\text{H}_4 +52$, $\text{C}_2\text{H}_6 -85$, $\text{H}_2 0$) to calculate ΔH_r . [3]

3.3

- $$\Delta H_r = \Delta H_f(\text{C}_2\text{H}_6) - [\Delta H_f(\text{C}_2\text{H}_4) + \Delta H_f(\text{H}_2)] = -85 - (52 + 0) = -137 \text{ kJ mol}^{-1}$$

3.4 Explain why Hess's law allows the enthalpy change of formation of a compound to be found even when the reaction cannot be carried out directly. [2]

3.4

- an indirect route (e.g. via combustion or formation of the elements) can be measured; by Hess's law the sum of the indirect steps equals the direct (unmeasurable) enthalpy change

Exercise 3.5 (a)

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
CH ₃ OH(l)	−239
CO ₂ (g)	−394
H ₂ O(l)	−286

(a) Write the equation for the complete combustion of methanol, then construct an energy cycle through the elements and use it to calculate $\Delta H_c^\ominus$ for methanol. [5]

Exercise 3.5 (a)

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
CH ₃ OH(l)	−239
CO ₂ (g)	−394
H ₂ O(l)	−286

(a) Write the equation for the complete combustion of methanol, then construct an energy cycle through the elements and use it to calculate $\Delta H_c^\ominus$ for methanol. [5]

(a)

- CH₃OH(l) + 1½O₂(g) → CO₂(g) + 2H₂O(l). The cycle runs from the elements up to the reactants and up to the products, so $\Delta H_c = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$; $= [(-394) + 2(-286)] - [(-239) + 0] = -966 + 239 = -727 \text{ kJ mol}^{-1}$

Exercise 3.5 (b)

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
CH ₃ OH(l)	−239
CO ₂ (g)	−394
H ₂ O(l)	−286

(b) Explain why $\Delta H_f^\ominus$ of O₂(g) is zero.

[2]

Exercise 3.5 (b)

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
CH ₃ OH(l)	−239
CO ₂ (g)	−394
H ₂ O(l)	−286

(b) Explain why $\Delta H_f^\ominus$ of O₂(g) is zero.

[2]

(b)

- By definition the standard enthalpy of formation of an **element in its standard state** is zero – forming it from itself is no change at all

Exercise 3.5 (c)

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
CH ₃ OH(l)	−239
CO ₂ (g)	−394
H ₂ O(l)	−286

(c) A school laboratory obtains -690 kJ mol^{-1} . State whether that is more or less exothermic than your answer, and give one reason for the difference. [2]

Exercise 3.5 (c)

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
CH ₃ OH(l)	−239
CO ₂ (g)	−394
H ₂ O(l)	−286

(c) A school laboratory obtains -690 kJ mol^{-1} . State whether that is more or less exothermic than your answer, and give one reason for the difference. [2]

(c)

- Less exothermic; heat escapes to the surroundings, so less energy appears to have been released per mole burned

Exercise 3.5 (d)

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
$\text{CH}_3\text{OH(l)}$	−239
$\text{CO}_2\text{(g)}$	−394
$\text{H}_2\text{O(l)}$	−286

(d) State the principle that allows this calculation even though methanol is never actually burned along the route you drew. [2]

Exercise 3.5 (d)

3.5 Use the standard enthalpy changes of formation below.

Compound	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
CH ₃ OH(l)	−239
CO ₂ (g)	−394
H ₂ O(l)	−286

(d) State the principle that allows this calculation even though methanol is never actually burned along the route you drew. [2]

(d)

- Hess's law: the enthalpy change of a reaction is independent of the route taken, provided the initial and final states are the same