

Exercise walk-through

2.1.2

Hess cycles and bond-enthalpy methods

Pearson Edexcel · International A-Level · Chemistry

Check the sign and denominator

Your turn

The solution gains 1.20 kJ while 0.0200 mol reacts. What is the molar reaction enthalpy?

- A. $+60.0 \text{ kJ mol}^{-1}$
- B. $-60.0 \text{ kJ mol}^{-1}$
- C. $-0.0240 \text{ kJ mol}^{-1}$
- D. $-1.20 \text{ kJ mol}^{-1}$

Supplied IR data: carbon environments

Bond / environment	Supplied region / cm^{-1}
Alkane C–H	2850–3000
Alkene C–H	3000–3100
Aldehyde C–H	2700–2900
C=C	1620–1680
C=O	1680–1750

Independent identification

Your turn

$\text{C}_3\text{H}_6\text{O}$ has a carbonyl IR absorption and gives a positive Fehling's test. Which structure fits all evidence?

- A. CH_3COCH_3
- B. $\text{CH}_3\text{CH}_2\text{CHO}$
- C. $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$
- D. CH_3COOH

Exercises 2.1·2.3

2.1 State what happens to ΔH when a reaction equation is reversed and when all coefficients are doubled. [2]

2.2 $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$ has $\Delta H = -184 \text{ kJ mol}^{-1}$. Calculate ΔH for $\text{HCl} \rightarrow \frac{1}{2}\text{H}_2 + \frac{1}{2}\text{Cl}_2$. [2]

2.3 Use $\text{H-H} = 436$, $\text{Cl-Cl} = 243$ and $\text{H-Cl} = 431 \text{ kJ mol}^{-1}$ to estimate ΔH for $\text{H}_2 +$

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2.1 Reversing changes the sign; doubling coefficients doubles ΔH .

2.2 $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$ has $\Delta H = -184 \text{ kJ mol}^{-1}$. Calculate ΔH for $\text{HCl} \rightarrow \frac{1}{2}\text{H}_2 + \frac{1}{2}\text{Cl}_2$. [2]

2.2 Reverse and halve: $+92 \text{ kJ mol}^{-1}$.

2.3 Use $\text{H-H} = 436$, $\text{Cl-Cl} = 243$ and $\text{H-Cl} = 431 \text{ kJ mol}^{-1}$ to estimate ΔH for $\text{H}_2 +$

2.3 Break one H-H and one Cl-Cl ; form two H-Cl . $\Delta H \approx 436 + 243 - 2(431) = -183 \text{ kJ mol}^{-1}$.

Exercises 3.1·3.2

3.1 Formation enthalpies are $\text{CH}_4(\text{g}) -75$, $\text{CO}_2(\text{g}) -394$ and $\text{H}_2\text{O}(\text{l}) -286 \text{ kJ mol}^{-1}$. Calculate ΔH for $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$. Explain the value used for oxygen. [4]

3.2 For $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g})$, $\Delta H = -72 \text{ kJ mol}^{-1}$. $\text{H-H} = 436$ and $\text{Br-Br} = 193 \text{ kJ mol}^{-1}$. Find the mean H-Br bond enthalpy and explain why these data cannot directly predict the rate. [4]

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3.1 $\Delta H = \sum \nu \Delta H_f(\text{products}) - \sum \nu \Delta H_f(\text{reactants})$. $[-394 + 2(-286)] - [-75] = -891$

3.2 For $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g})$, $\Delta H = -72 \text{ kJ mol}^{-1}$. $\text{H-H} = 436$ and $\text{Br-Br} = 193 \text{ kJ mol}^{-1}$. Find the mean H-Br bond enthalpy and explain why these data cannot directly predict the rate. [4]

3.2 $-72 = 436 + 193 - 2E(\text{H-Br})$; $E = (436 + 193 + 72)/2 = 350.5 \text{ kJ mol}^{-1}$. Rate depends on the activation pathway, not just net enthalpy or a bond sum.

Exercise 3.3

3.3 Use the following measured enthalpies to find ΔH for $\text{Mg(s)} + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{MgO(s)}$:
 $\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})$, -466 kJ mol^{-1} ; $\text{MgO(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2\text{O(l)}$, -146 kJ mol^{-1} ; $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O(l)}$, -286 kJ mol^{-1} . Show which equation is reversed and one experimental limitation. [5]

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3.3 Reverse the MgO reaction, giving $+146 \text{ kJ mol}^{-1}$. Add the Mg reaction and hydrogen combustion, cancel HCl, MgCl_2 , H_2 and H_2O to obtain the target. $\Delta H = -466 - 286 + 146 = -606 \text{ kJ mol}^{-1}$. Heat loss in the measured reactions or unmatched solution states/concentrations can limit the experimental cycle.