

2.1.2 Hess cycles and bond-enthalpy methods

Name: _____ Class: _____ Date: _____

Total: 20 marks

KEY WORDS Hess's law

Objective

Choose and check a thermochemical route, including reversal, coefficients and gaseous-bond assumptions. All questions and marking guidance are original.

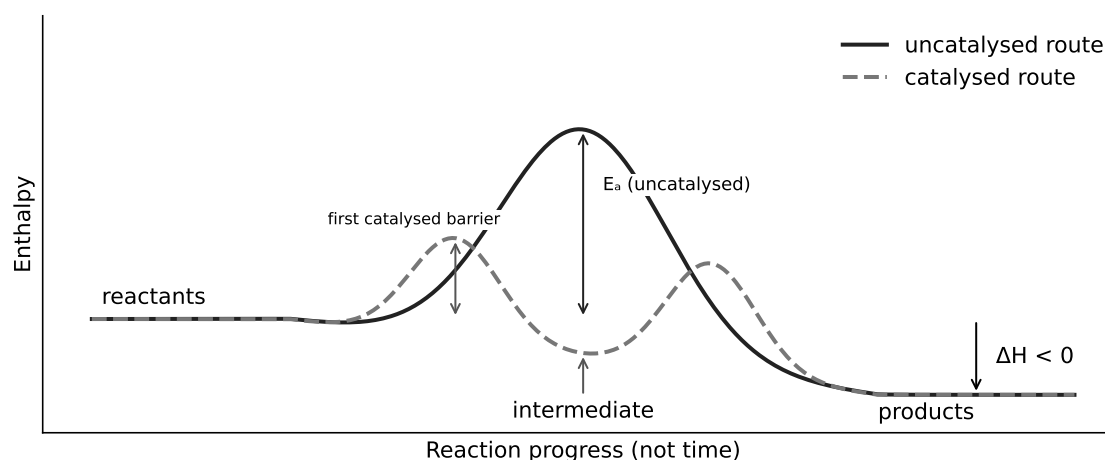
1 Key knowledge and worked examples

Hess's law compares routes with identical initial and final states. Reverse a chemical equation and its enthalpy sign; multiply both together. Cancel species before summing energy changes.

■ Worked method

For $\text{C(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{CO(g)}$, use $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$, -394 kJ mol^{-1} , and reverse $\text{CO} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}_2$, -283 kJ mol^{-1} . Cancel CO_2 and $\frac{1}{2}\text{O}_2$ to get the target. $\Delta H = -394 - (-283) = -111 \text{ kJ mol}^{-1}$.

For a gas-phase bond estimate, count every bond on each side, including coefficients: $\Delta H \approx \sum E(\text{broken}) - \sum E(\text{formed})$. A mean bond enthalpy averages different molecules, so it is not an exact value for every environment.



2 Practice

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 $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$. [3]

3 Exam-style questions

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]

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

Solutions

Original marking guidance: M1 method, A1 accuracy, B1 explanation.

2.1 Reversing changes the sign; doubling coefficients doubles ΔH .

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

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

3.1 $\Delta H = \sum \nu \Delta H_f(\text{products}) - \sum \nu \Delta H_f(\text{reactants})$. $[-394 + 2(-286)] - [-75] = -891 \text{ kJ mol}^{-1}$. Oxygen has zero formation enthalpy in its standard state by convention.

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

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.