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.

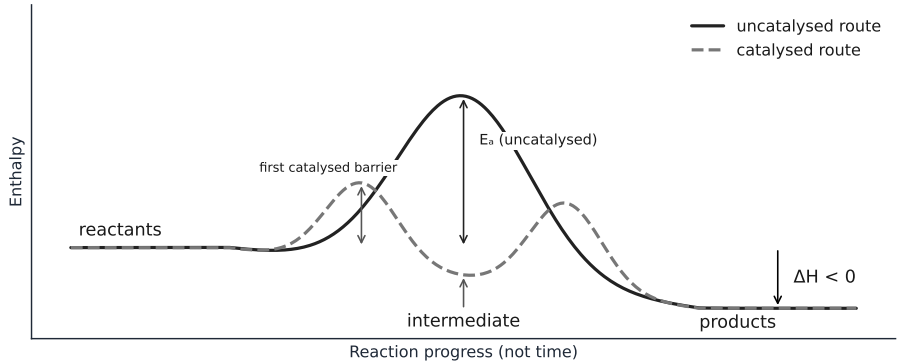
1 Key knowledge and worked examples

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 $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$, use $C + O_2 \rightarrow CO_2$, -394 kJ mol^{-1}, and reverse $CO + \frac{1}{2}O_2 \rightarrow CO_2$, -283 kJ mol^{-1}. Cancel CO_2 and $\frac{1}{2}O_2$ to get the target. $\Delta H = -394 - (-283) = -111$ 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 $H_2 + Cl_2 \rightarrow 2HCl$ has $\Delta H = -184$ kJ mol^{-1}. Calculate ΔH for $HCl \rightarrow \frac{1}{2}H_2 + \frac{1}{2}Cl_2$. [2]

2.3 Use H–H = 436, Cl–Cl = 243 and H–Cl = 431 kJ mol^{-1} to estimate ΔH for $H_2 + Cl_2 \rightarrow 2HCl$. [3]

3 Exam-style questions

3.1 Formation enthalpies are $CH_4(g)$ −75, $CO_2(g)$ −394 and $H_2O(l)$ −286 kJ mol^{-1}. Calculate ΔH for $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$. Explain the value used for oxygen. [4]

3.2 For $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$, $\Delta H = -72$ kJ mol^{-1}. H–H = 436 and Br–Br = 193 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 $Mg(s) + \tfrac{1}{2}O_2(g) \rightarrow MgO(s)$: $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$, -466 kJ mol^{-1}; $MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$, -146 kJ mol^{-1}; $H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)$, -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 kJ mol^{-1}.

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

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

3.2 $-72 = 436 + 193 - 2E(\text{H} - \text{Br})$; $E = (436 + 193 + 72)/2 = 350.5$ 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 kJ mol^{-1}. Add the Mg reaction and hydrogen combustion, cancel HCl, $MgCl_2$, H_2 and H_2O to obtain the target. ΔH = $-466 - 286 + 146 = -606$ kJ mol^{-1}. Heat loss in the measured reactions or unmatched solution states/concentrations can limit the experimental cycle.