

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.