Solutions

Each answer shows the steps, then the **result**.

2.1

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

2.2

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

3.1 B

- Hess's law follows from the conservation of energy

3.2 B

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

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

- 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

3.5

(a)

- $\text{CH}_3\text{OH(l)} + 1\frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + 2\text{H}_2\text{O(l)}$. The cycle runs from the elements up to the reactants and up to the products, so $\Delta H_\text{c} = \sum \Delta H_\text{f}(\text{products}) - \sum \Delta H_\text{f}(\text{reactants})$; $= [(-394) + 2(-286)] - [(-239) + 0] = -966 + 239 = -727 \text{ kJ mol}^{-1}$

(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

(c)

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

(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