Hess's Law

SUMMARY Enthalpy of Reaction and Hess's Law

To determine an enthalpy change of a reaction by using Hess's law, follow these steps:

- 1. Write the net reaction equation if it is not given.
- 2. Manipulate the known equations so they will add to yield the target equation.
- 3. Cancel and add the remaining reactants and products.
- 4. Add the component enthalpy changes to obtain the net enthalpy change.
- 5. Determine the molar enthalpy, if required.

Example

What is the enthalpy change for the formation of one mole of butane (C_4H_{10}) gas from its elements? The reaction is:

 $4 C_{(s)} + 5 H_{2(g)} \rightarrow C_4 H_{10(g)} \qquad \Delta H^\circ = ?$

The following known equations, determined by calorimetry, are provided:

(1)
$$C_4 H_{10(g)} + \frac{15}{2} O_{2(g)} \rightarrow 4 CO_{2(g)} + 5 H_2 O_{(g)} \Delta H_1^\circ = -2657.4 \text{ kJ}$$

(2) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} \Delta H_2^\circ = -393.5 \text{ kJ}$

(3) 2 H_{2(g)} + O_{2(g)}
$$\rightarrow$$
 2 H₂O_(g) $\Delta H_3^{\circ} = -483.6 \text{ kJ}$

Reversing known equation (1), which will require multiplying its ΔH by -1, will make $C_4H_{10(g)}$ a product; multiplying known equation (2) by 4 will provide the required amount of $C_{(s)}$ reactant; and multiplying known equation (3) by 5/2 will provide the required amount of $H_{2(g)}$ reactant. Cancellation when the equations are added will determine whether the required amount of $O_{2(g)}$ remains.

$$-1 \times (1): \quad 4 \operatorname{CO}_{2(g)} + 5 \operatorname{H}_2 \operatorname{O}_{(g)} \quad \rightarrow \operatorname{C}_4 \operatorname{H}_{10(g)} + \frac{13}{2} \operatorname{O}_{2(g)}$$

$$\Delta H^{\circ} = -1(-657.4) \text{ kJ}$$

$$4 \times (2):$$
 $4 C_{(s)} + 4 O_{2(g)} \rightarrow 4 CO_{2(g)} \Delta H^{\circ} = 4(-393.5) \text{ kJ}$

$$\frac{5}{2} \times (3):$$
 5 H_{2(g)} + $\frac{5}{2}$ O_{2(g)} \rightarrow 5 H₂O_(g) $\Delta H^{\circ} = \frac{5}{2}(-483.6)$ kJ

$$\begin{array}{rll} 4 \ \mathcal{GO}_{2(g)} \ + \ 5 \ \mathcal{H}_2 \dot{\mathcal{O}}_{(g)} \ + \ 4 \ \mathcal{C}_{(s)} \ + \ \frac{13}{2} \ \mathcal{O}_{2(g)} \ + \ 5 \ \mathcal{H}_{2(g)} \rightarrow \ \mathcal{C}_4 \mathcal{H}_{10(g)} \ + \ \frac{13}{2} \ \mathcal{O}_{2(g)} \\ & + \ 4 \ \mathcal{GO}_{2(g)} \ + \ 5 \ \mathcal{H}_2 \dot{\mathcal{O}}_{(g)} \\ & \text{or} & 4 \ \mathcal{C}_{(s)} \ + \ 5 \ \mathcal{H}_{2(g)} & \longrightarrow \ \mathcal{C}_4 \mathcal{H}_{10(g)} \end{array}$$

(continued)

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If the known equations can be added together to form the target equation, *then* their enthalpy changes can be added together. In this case,

$$\Delta H^{\circ}_{\text{total}} = (+2657.4) + (-1574.0) + (-1209.0) \text{ kJ}$$

$$\Delta H^{\circ}_{\text{total}} = -125.6 \text{ kJ}$$

The enthalpy change for the formation of one mole of butane is -125.6 kJ.