Molar Enthalpies of Formation

SUMMARY Using Enthalpies of Formation to Predict ΔH

According to Hess's law, the net enthalpy change for a chemical reaction is equal to the sum of the enthalpies of formation of the products minus the sum of the enthalpies of formation of the reactants.

$$\Delta H = \Sigma n H_{f(\text{products})} - \Sigma n H_{f(\text{reactants})}$$

Example

What is the standard molar enthalpy for the slaking of lime?

$\mathbf{CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(s)}}$ $\Delta H° = ?$

$$\Delta H = \sum n\Delta H°_{f(\text{products})} - \sum n\Delta H°_{f(\text{reactants})}$$

where n represents the amount (in moles) of each particular product or reactant. Substitute our known molar enthalpies of formation into this equation:

$$\Delta H = n\Delta H°_{f(Ca(OH)_{(s)})} - (n\Delta H°_{f(CaO_{(s)})} + n\Delta H°_{f(H_2O)_{(l)}})$$

$$= -\left(1 \text{ mol} \times \frac{986.1 \text{ kJ}}{1 \text{ mol}}\right) - \left((1 \text{ mol} \times \frac{-634.9 \text{ kJ}}{1 \text{ mol}}) + (1 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}}\right)$$

$$\Delta H = -65.4 \text{ kJ}$$

$$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(s)} \qquad \Delta H = -65.4 \text{ kJ}$$

Example

What is the standard molar enthalpy of combustion of methane fuel?

$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(l)}$

$$\Delta H = \sum n\Delta H°_{f(\text{products})} - \sum n\Delta H°_{f(\text{reactants})}$$

$$= \left(1 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}}) + 2 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}}\right) -$$

$$(1 \text{ mol} \times \frac{-74.4 \text{ kJ}}{1 \text{ mol}} + 2 \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}})$$

$$= -965.1 \text{ kJ} - (-74.4 \text{ kJ})$$

$$\Delta H = -890.7 \text{ kJ}$$

$$\Delta H_c = \frac{\Delta H}{n} = \frac{-890.7 \text{ kJ}}{1 \text{ mol } CH_4}$$

$$= -890.7 \text{ kJ/mol } CH_4$$

The molar enthalpy of combustion of methane fuel is –890.7 kJ/mol.