To find the necessary sums, the ΔH_f^0 value for each reactant and each product must be multiplied by its respective coefficient in the desired equation. For the reaction of NO with O_2, applying this equation gives the following value for ΔH^0.

$$\begin{aligned}\Delta H^0 &= \Delta H_f^0(NO_2) - [\Delta H_f^0(NO) + 0]\\ &= +33.2\ \text{kJ/mol} - 90.29\ \text{kJ/mol} = -57.1\ \text{kJ}\end{aligned}$$

Note that zero is the assigned value for the enthalpies of formation of elements in their standard states.

PRACTICE

Answers in Appendix E

1. Calculate the enthalpy of reaction for the combustion of methane gas, CH_4, to form $CO_2(g) + H_2O(l)$.
2. Carbon occurs in two distinct forms. It can be the soft, black material found in pencils and lock lubricants, called graphite, or it can be the hard, brilliant gem we know as diamond. Calculate ΔH^0 for the conversion of graphite to diamond for the following reaction.

$$C_{graphite}(s) \longrightarrow C_{diamond}(s)$$

The combustion reactions you will need follow.

$$C_{graphite}(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_c^0 = -394\ \text{kJ}$$
$$C_{diamond}(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_c^0 = -396\ \text{kJ}$$

extension

Go to **go.hrw.com** for more practice problems that ask you to use Hess's law.

Keyword: HC6NRGX

Determining Enthalpy of Formation

When carbon is burned in a limited supply of oxygen, carbon monoxide is produced. In this reaction, carbon is probably first oxidized to carbon dioxide. Then part of the carbon dioxide is reduced with carbon to give some carbon monoxide. Because these two reactions occur simultaneously and we get a mixture of CO and CO_2, it is not possible to directly measure the enthalpy of formation of $CO(g)$ from $C(s)$ and $O_2(g)$.

$$C(s) + \tfrac{1}{2}O_2(g) \longrightarrow CO(g) \qquad \Delta H_f^0 = ?$$

However, we do know the enthalpy of formation of carbon dioxide and the enthalpy of combustion of carbon monoxide.

$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_f^0 = -393.5\ \text{kJ/mol}$$
$$CO(g) + \tfrac{1}{2}O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_c^0 = -283.0\ \text{kJ/mol}$$

We reverse the second equation because we need CO as a product. Adding gives the desired enthalpy of formation of carbon monoxide.