



## 9.10: Determining Enthalpies of Reaction from Standard Enthalpies of Formation

We have examined how to measure  $\Delta H$  for a chemical reaction directly through calorimetry and how to infer  $\Delta H$  for a reaction from related reactions through Hess's law. We have also seen how to estimate  $\Delta H$  for a reaction from bond energies. We now turn to a fourth and more convenient way to determine  $\Delta H$  for a large number of chemical reactions: from tabulated *standard enthalpies of formation*.

### Standard States and Standard Enthalpy Changes

Recall that  $\Delta H$  is the *change* in enthalpy for a chemical reaction—the difference in enthalpy between the products and the reactants. Because we are interested in changes in enthalpy (and not in absolute values of enthalpy itself), we are free to define the *zero* of enthalpy as conveniently as possible. Returning to our mountain-climbing analogy from [Section 9.3](#), a change in altitude (like a change in enthalpy) is an absolute quantity. Altitude itself (like enthalpy), however, is a relative quantity, defined relative to some standard (such as sea level in the case of altitude). We define a similar, albeit slightly more complex, standard for enthalpy. This standard has three parts: the standard state, the standard enthalpy change ( $\Delta H^\circ$ ), and the standard enthalpy (or heat) of formation ( $\Delta H_f^\circ$ ).

#### 1. Standard State

- *For a Gas:* The standard state for a gas is the pure gas at a pressure of exactly 1 atm.
- *For a Liquid or Solid:* The standard state for a liquid or solid is the pure substance in its most stable form at a pressure of 1 atm and at the temperature of interest (often taken to be 25 °C).
- *For a Substance in Solution:* The standard state for a substance in solution is a concentration of exactly 1 M.

#### 2. Standard Enthalpy Change ( $\Delta H^\circ$ )

- The change in enthalpy for a process when all reactants and products are in their standard states. The degree sign indicates standard states.

#### 3. Standard Enthalpy of Formation ( $\Delta H_f^\circ$ )

- *For a Pure Compound:* The change in enthalpy when 1 mol of the compound forms from its constituent elements in their standard states.
- *For a Pure Element in Its Standard State:*  $\Delta H_f^\circ = 0$

The standard state was changed in 1997 to a pressure of 1 bar, which is very close to 1 atm (1 atm = 1.013 bar). Both standards are now in common use.

The standard enthalpy of formation is also called the standard heat of formation.

Assigning the value of zero to the standard enthalpy of formation for an element in its standard state is the equivalent of assigning an altitude of zero to sea level. Once we assume sea level is zero, we can measure all subsequent changes in altitude relative to sea level. Similarly, we can measure all changes in enthalpy relative to those of pure elements in their standard states.

For example, consider the standard enthalpy of formation of methane gas at 25 °C:



The carbon in this equation must be graphite (the most stable form of carbon at 1 atm and 25 °C).

For methane, as with most compounds,  $\Delta H_f^\circ$  is negative. Continuing our analogy, if we think of pure elements in their standard states as being at *sea level*, then most compounds lie *below sea level*. We always write the chemical equation for the enthalpy of formation of a compound to form 1 mol of the compound, so  $\Delta H_f^\circ$  has the units of kJ/mol. [Table 9.4](#) lists  $\Delta H_f^\circ$  values for selected compounds. A more comprehensive list is in [Appendix IIB](#).

**Table 9.4 Standard Enthalpies (or Heats) of Formation,  $\Delta H_f^\circ$ , at 298 K**

Formula	$\Delta H_f^\circ$ (kJ/mol)	Formula	$\Delta H_f^\circ$ (kJ/mol)	Formula
<b>Bromine</b>		$\text{C}_3\text{H}_8\text{O}(\text{l, isopropanol})$	-318.1	<b>Oxygen</b>
$\text{Br}(\text{g})$	111.9	$\text{C}_6\text{H}_6(\text{l})$	49.1	$\text{O}_2(\text{g})$
$\text{Br}_2(\text{l})$	0	$\text{C}_6\text{H}_{12}\text{O}_6(\text{s, glucose})$	-1273.3	$\text{O}_3(\text{g})$
$\text{HBr}(\text{g})$	-36.3	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s, sucrose})$	-2226.1	$\text{H}_2\text{O}(\text{g})$
<b>Calcium</b>		<b>Chlorine</b>		$\text{H}_2\text{O}(\text{l})$
$\text{Ca}(\text{s})$	0	$\text{Cl}(\text{g})$	121.3	<b>Silver</b>
$\text{CaO}(\text{s})$	-634.9	$\text{Cl}_2(\text{g})$	0	$\text{Ag}(\text{s})$
$\text{CaCO}_3(\text{s})$	-1207.6	$\text{HCl}(\text{g})$	-92.3	$\text{AgCl}(\text{s})$
<b>Carbon</b>		<b>Fluorine</b>		<b>Sodium</b>
$\text{C}(\text{s, graphite})$	0	$\text{F}(\text{g})$	79.38	$\text{Na}(\text{s})$
$\text{C}(\text{s, diamond})$	1.88	$\text{F}_2(\text{g})$	0	$\text{Na}(\text{g})$
$\text{CO}(\text{g})$	-110.5	$\text{HF}(\text{g})$	-273.3	$\text{NaCl}(\text{s})$
$\text{CO}_2(\text{g})$	-393.5	<b>Hydrogen</b>		$\text{Na}_2\text{CO}_3(\text{s})$
$\text{CH}_4(\text{g})$	-74.6	$\text{H}(\text{g})$	218.0	$\text{NaHCO}_3(\text{s})$
$\text{CH}_3\text{OH}(\text{l})$	-238.6	$\text{H}_2(\text{g})$	0	<b>Sulfur</b>
$\text{C}_2\text{H}_2(\text{g})$	227.4	<b>Nitrogen</b>		$\text{S}_8(\text{s, rhombic})$
$\text{C}_2\text{H}_4(\text{g})$	52.4	$\text{N}_2(\text{g})$	0	$\text{S}_8(\text{s, monoclinic})$
$\text{C}_2\text{H}_6(\text{g})$	-84.68	$\text{NH}_3(\text{g})$	-45.9	$\text{SO}_2(\text{g})$
$\text{C}_2\text{H}_5\text{OH}(\text{l})$	-277.6	$\text{NH}_4\text{NO}_3(\text{s})$	-365.6	$\text{SO}_3(\text{g})$
$\text{C}_3\text{H}_8(\text{g})$	-103.85	$\text{NO}(\text{g})$	91.3	$\text{H}_2\text{SO}_4(\text{l})$
$\text{C}_3\text{H}_6\text{O}(\text{l, acetone})$	-248.4	$\text{N}_2\text{O}(\text{g})$	81.6	

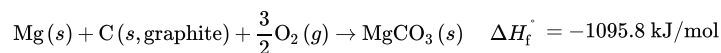
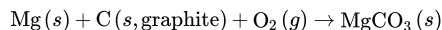
### Example 9.11 Standard Enthalpies of Formation

Write equations for the formation of (a)  $\text{MgCO}_3(\text{s})$  and (b)  $\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$  from their respective elements in their standard states. Include the value of  $\Delta H_f^\circ$  for each equation.

**SOLUTION**(a)  $\text{MgCO}_3(s)$ 

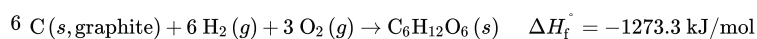
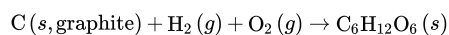
Write the equation with the constituent elements in  $\text{MgCO}_3$  in their standard states as the reactants and 1 mol of  $\text{MgCO}_3$  as the product.

Balance the equation and look up  $\Delta H_f^\circ$  in **Appendix IIB**. (Use fractional coefficients so that the product of the reaction is 1 mol of  $\text{MgCO}_3$ .)

(b)  $\text{C}_6\text{H}_{12}\text{O}_6(s)$ 

Write the equation with the constituent elements in  $\text{C}_6\text{H}_{12}\text{O}_6$  in their standard states as the reactants and 1 mol of  $\text{C}_6\text{H}_{12}\text{O}_6$  as the product.

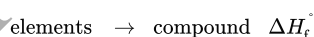
Balance the equation and look up  $\Delta H_f^\circ$  in **Appendix IIB**.



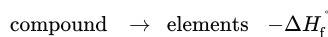
**FOR PRACTICE 9.11** Write equations for the formation of (a)  $\text{NaCl}(s)$  and (b)  $\text{Pb}(\text{NO}_3)_2(s)$  from their respective elements in their standard states. Include the value of  $\Delta H_f^\circ$  for each equation.

## Calculating the Standard Enthalpy Change for a Reaction

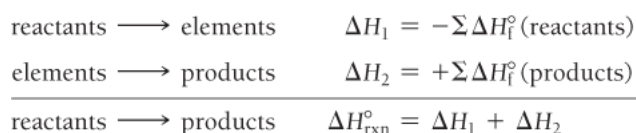
We have just seen that the standard enthalpy of formation corresponds to the *formation* of a compound from its constituent elements in their standard states:



Therefore, the *negative* of the standard enthalpy of formation corresponds to the *decomposition* of a compound into its constituent elements in their standard states:

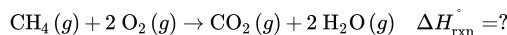


We can use these two concepts—the decomposition of a compound into its elements and the formation of a compound from its elements—to calculate the enthalpy change of any reaction by mentally taking the reactants through two steps. In the first step we *decompose the reactants* into their constituent elements in their standard states; in the second step we *form the products* from the constituent elements in their standard states:



In these equations,  $\Sigma$  means “the sum of” so that  $\Delta H_1$  is the sum of the negatives of the heats of formation of the reactants and  $\Delta H_2$  is the sum of the heats of formation of the products.

We can demonstrate this procedure by calculating the standard enthalpy change ( $\Delta H_{\text{rxn}}^\circ$ ) for the combustion of methane:



The enthalpy changes associated with the decomposition of the reactants and the formation of the products are shown in Figure 9.11. The first step (1) is the decomposition of 1 mol of methane into its constituent elements in their standard states. We can obtain the change in enthalpy for this step by reversing the enthalpy of formation equation for methane and changing the sign of  $\Delta H_f^\circ$ :

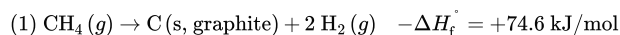
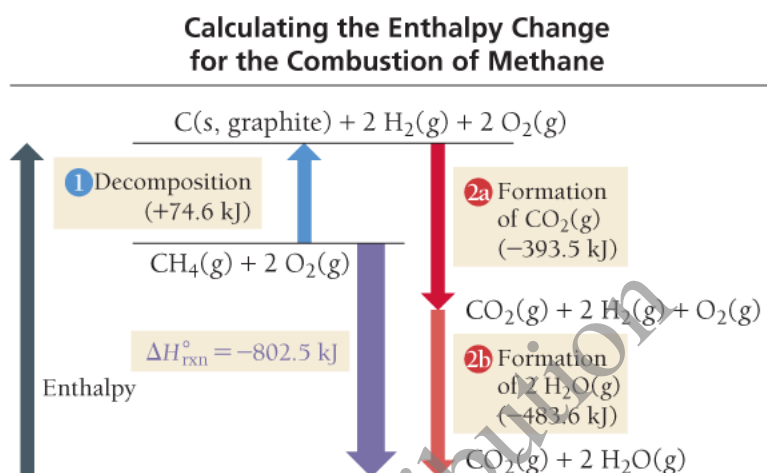
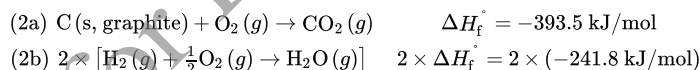


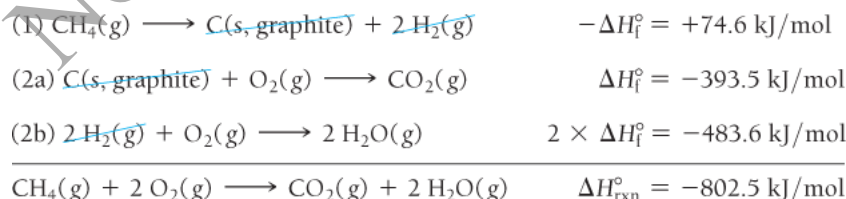
Figure 9.11 Calculating the Enthalpy Change for the Combustion of Methane



The second step, the formation of the products from their constituent elements, has two parts: (2a) the formation of 1 mol  $\text{CO}_2$  and (2b) the formation of 2 mol  $\text{H}_2\text{O}$ . Because part (2b) forms 2 mol  $\text{H}_2\text{O}$ , we multiply the  $\Delta H_f^\circ$  for that step by 2:



As we know from Hess's law, the enthalpy of reaction for the overall reaction is the sum of the enthalpies of reaction of the individual steps:



We can streamline and generalize this process as follows:

To calculate  $\Delta H_{\text{rxn}}^\circ$ , subtract the enthalpies of formation of the reactants multiplied by their stoichiometric coefficients from the enthalpies of formation of the products multiplied by their stoichiometric coefficients.

In the form of an equation,

[9.14]

$$\Delta H_{\text{rxn}}^{\circ} = \sum n_p \Delta H_f^{\circ} (\text{products}) - \sum n_r \Delta H_f^{\circ} (\text{reactants})$$

In this equation,  $n_p$  represents the stoichiometric coefficients of the products,  $n_r$  represents the stoichiometric coefficients of the reactants, and  $\Delta H_f^{\circ}$  represents the standard enthalpies of formation. Keep in mind when using this equation that elements in their standard states have  $\Delta H_f^{\circ} = 0$  **Examples 9.12** and **9.13** demonstrate this process.

### Example 9.12 $\Delta H_{\text{rxn}}^{\circ}$ and Standard Enthalpies of Formation

Use the standard enthalpies of formation to determine  $\Delta H_{\text{rxn}}^{\circ}$  for the reaction.



**SORT** You are given the balanced equation and asked to find the enthalpy of reaction.

**GIVEN:**  $4 \text{NH}_3(g) + 5 \text{O}_2(g) \rightarrow 4 \text{NO}(g) + 6 \text{H}_2\text{O}(g)$

**FIND:**  $\Delta H_{\text{rxn}}^{\circ}$

**STRATEGIZE** To calculate  $\Delta H_{\text{rxn}}^{\circ}$  from standard enthalpies of formation, subtract the heats of formation of the reactants multiplied by their stoichiometric coefficients from the heats of formation of the products multiplied by their stoichiometric coefficients.

**CONCEPTUAL PLAN**

$$\Delta H_{\text{rxn}}^{\circ} = \sum n_p \Delta H_f^{\circ} (\text{products}) - \sum n_r \Delta H_f^{\circ} (\text{reactants})$$

**SOLVE** Begin by looking up (in **Appendix IIB**) the standard enthalpy of formation for each reactant and product. Remember that the standard enthalpy of formation of pure elements in their standard state is zero. Calculate  $\Delta H_{\text{rxn}}^{\circ}$  by substituting into the equation.

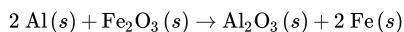
#### SOLUTION

Reactant or product	$\Delta H_f^{\circ} (\text{kJ/mol, from Appendix IIB})$
$\text{NH}_3(g)$	-45.9
$\text{O}_2(g)$	0.0
$\text{NO}(g)$	+91.3
$\text{H}_2\text{O}(g)$	-241.8

$$\begin{aligned}
 \Delta H_{\text{rxn}}^{\circ} &= \sum n_p \Delta H_f^{\circ} (\text{products}) - \sum n_r \Delta H_f^{\circ} (\text{reactants}) \\
 &= [4(\Delta H_{f, \text{NO}(g)}^{\circ}) + 6(\Delta H_{f, \text{H}_2\text{O}(g)}^{\circ})] - [4(\Delta H_{f, \text{NH}_3(g)}^{\circ}) + 5(\Delta H_{f, \text{O}_2(g)}^{\circ})] \\
 &= [4(+91.3 \text{ kJ}) + 6(-241.8 \text{ kJ})] - [4(-45.9 \text{ kJ}) + 5(0.0 \text{ kJ})] \\
 &= -1085.6 \text{ kJ} - (-183.6 \text{ kJ}) \\
 &= -902.0 \text{ kJ}
 \end{aligned}$$

**CHECK** The units of the answer (kJ) are correct. The answer is negative, which means that the reaction is exothermic.

**FOR PRACTICE 9.12** The thermite reaction, in which powdered aluminum reacts with iron(III) oxide, is highly exothermic.



Use standard enthalpies of formation to find  $\Delta H_{\text{rxn}}^\circ$  for the thermite reaction.



The reaction of powdered aluminum with iron(III) oxide, known as the thermite reaction, releases a large amount of heat.

### Interactive Worked Example 9.12 $\Delta H_{\text{rxn}}^\circ$ and Standard Enthalpies of Formation

### Example 9.13 $\Delta H_{\text{rxn}}^\circ$ and Standard Enthalpies of Formation

A city of 100,000 people uses approximately  $1.0 \times 10^{11}$  kJ of energy per day. Suppose all of the city's energy comes from the combustion of liquid octane ( $\text{C}_8\text{H}_{18}$ ) to form gaseous water and gaseous carbon dioxide. Use standard enthalpies of formation to calculate  $\Delta H_{\text{rxn}}^\circ$  for the combustion of octane and determine how many kilograms of octane are necessary to provide this amount of energy.

**SORT** You are given the amount of energy used and asked to find the mass of octane required to produce the energy.

**GIVEN:**  $1.0 \times 10^{11}$  kJ

**FIND:** kg  $\text{C}_8\text{H}_{18}$

**STRATEGIZE** The conceptual plan has three parts. In the first part, write a balanced equation for the combustion of octane.

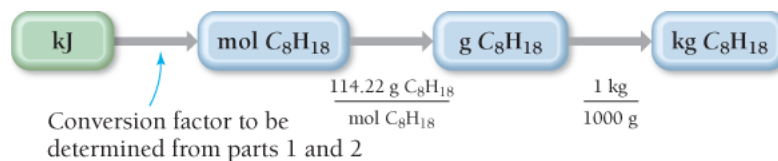
In the second part, calculate  $\Delta H_{\text{rxn}}^\circ$  from the  $\Delta H_f^\circ$  values of the reactants and products.

In the third part, convert from kilojoules of energy to moles of octane using the conversion factor found in part 2, and then convert from moles of octane to mass of octane using the molar mass.

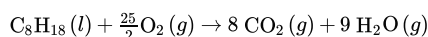
#### CONCEPTUAL PLAN

1. Write balanced equation.
- 2.



**RELATIONSHIPS USED**molar mass  $\text{C}_8\text{H}_{18} = 114.22 \text{ g/mol}$  $1 \text{ kg} = 1000 \text{ g}$ 

**SOLVE** Begin by writing the balanced equation for the combustion of octane. For convenience, do not clear the  $\frac{25}{2}$  fraction in order to keep the coefficient on octane as 1.

**SOLUTION PART 1**

Look up (in **Appendix IIB**) the standard enthalpy of formation for each reactant and product and calculate  $\Delta H_{\text{rxn}}^\circ$ .

**SOLUTION PART 2**

Reactant or product	$\Delta H_f^\circ$ (kJ/mol, from Appendix IIB)
$\text{C}_8\text{H}_{18}(l)$	-250.1
$\text{O}_2(g)$	0.0
$\text{CO}_2(g)$	-393.5
$\text{H}_2\text{O}(g)$	-241.8

$$\begin{aligned}
 \Delta H_{\text{rxn}}^\circ &= \sum n_p \Delta H_f^\circ (\text{products}) - \sum n_r \Delta H_f^\circ (\text{reactants}) \\
 &= [8(\Delta H_f^\circ, \text{CO}_2(g)) + 9(\Delta H_f^\circ, \text{H}_2\text{O}(g))] - [1(\Delta H_f^\circ, \text{C}_8\text{H}_{18}(l)) + \frac{25}{2}(\Delta H_f^\circ, \text{O}_2(g))] \\
 &= [8(-393.5 \text{ kJ}) + 9(-241.8 \text{ kJ})] - [1(-250.1 \text{ kJ}) + \frac{25}{2}(0.0 \text{ kJ})] \\
 &= -5324.2 \text{ kJ} - (-250.1 \text{ kJ}) \\
 &= -5074.1 \text{ kJ}
 \end{aligned}$$

From parts 1 and 2 of the conceptual plan, build a conversion factor between mol  $\text{C}_8\text{H}_{18}$  and kJ.

Follow part 3 of the conceptual plan. Begin with  $-1.0 \times 10^{11} \text{ kJ}$  (because the city uses this much energy, the reaction must emit this amount, and therefore the sign is negative) and follow the steps to determine kg octane.

**SOLUTION PART 3**

$$1 \text{ mol C}_8\text{H}_{18} : -5074.1 \text{ kJ}$$

$$-1.0 \times 10^{11} \text{ kJ} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{-5074.1 \text{ kJ}} \times \frac{114.22 \text{ g C}_8\text{H}_{18}}{1 \text{ mol C}_8\text{H}_{18}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 2.3 \times 10^6 \text{ kg C}_8\text{H}_{18}$$

**CHECK** The units of the answer (kg  $\text{C}_8\text{H}_{18}$ ) are correct. The answer is positive, as it should be for

mass. The magnitude is fairly large, as you would expect because this amount of octane is supposed to provide the energy for an entire city.

**FOR PRACTICE 9.13** Chemical hand warmers popular with skiers and snowboarders produce heat when they are removed from their airtight plastic wrappers. They utilize the oxidation of iron to form iron(III) oxide according to the reaction:  $4 \text{Fe}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Fe}_2\text{O}_3(s)$ . Calculate  $\Delta H_{\text{rxn}}^\circ$  for this reaction and calculate how much heat is produced from a hand warmer containing 15.0 g of iron powder.

Not for Distribution



*Not for Distribution*