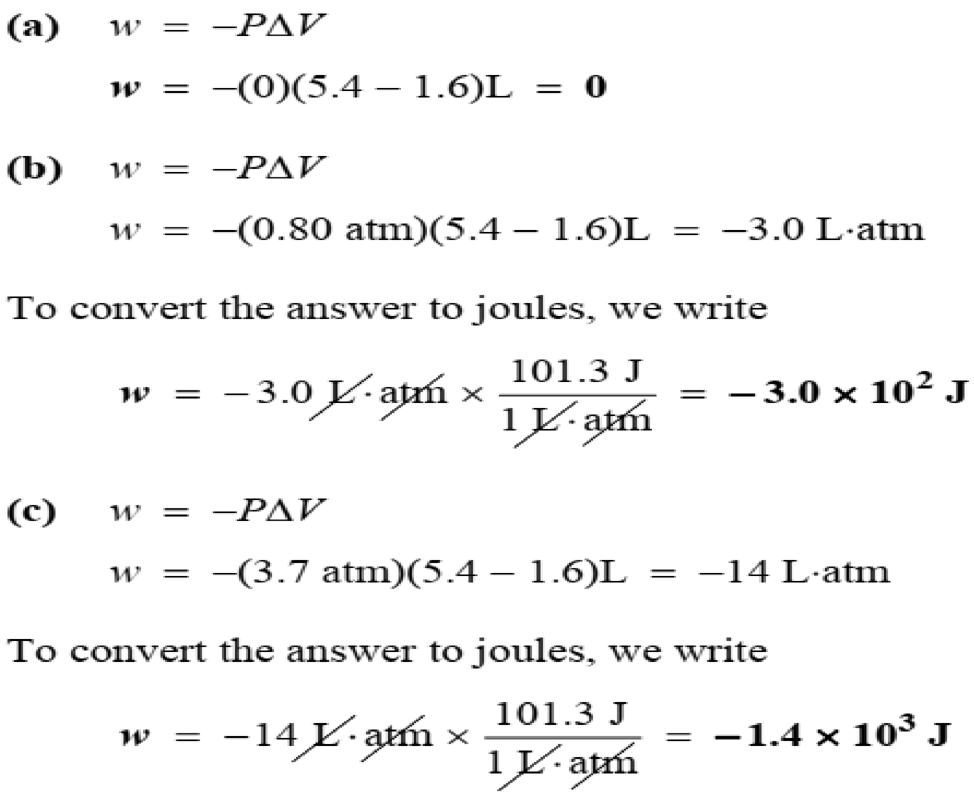
CHEM 101

Chapter 6 Problems

6.16 A sample of nitrogen gas expands in volume from 1.6 L to 5.4 L at constant temperature. Calculate the work done in joules if the gas expands (a) against a vacuum, (b) against a constant pressure of 0.80 atm, and (c) against a constant pressure of 3.7 atm



(a) $w = -P\Delta V$

$\boldsymbol{w} = -(0)(5.4 - 1.6)\text{L} = \mathbf{0}$

(b) $w = -P\Delta V$

$w = -(0.80\ \text{atm})(5.4 - 1.6)\text{L} = -3.0\ \text{L}\cdot\text{atm}$

To convert the answer to joules, we write

$$\boldsymbol{w} = -3.0\ \cancel{\text{L}}\cdot\cancel{\text{atm}} \times \frac{101.3\ \text{J}}{1\ \cancel{\text{L}}\cdot\cancel{\text{atm}}} = \mathbf{-3.0 \times 10^2\ J}$$

(c) $w = -P\Delta V$

$w = -(3.7\ \text{atm})(5.4 - 1.6)\text{L} = -14\ \text{L}\cdot\text{atm}$

To convert the answer to joules, we write

$$\boldsymbol{w} = -14\ \cancel{\text{L}}\cdot\cancel{\text{atm}} \times \frac{101.3\ \text{J}}{1\ \cancel{\text{L}}\cdot\cancel{\text{atm}}} = \mathbf{-1.4 \times 10^3\ J}$$

6.17 A gas expands and does *P-V* work on the surroundings equal to 325 J. At the same time, it absorbs 127 J of heat from the surroundings. Calculate the change in energy of the gas

An expansion implies an increase in volume, therefore w must be -325 J. If the system absorbs heat, q must be +127 J. The change in energy (internal energy) is:
$\Delta \boldsymbol{E} = q + w = 127\ \text{J} - 325\ \text{J} = \mathbf{-198\ J}$

6.18 The work done to compress a gas is 74 J. As a result, 26 J of heat is given off to the surroundings. Calculate the change in energy of the gas

Work of compression is positive.

Heat is given off by the gas, *q* is negative.

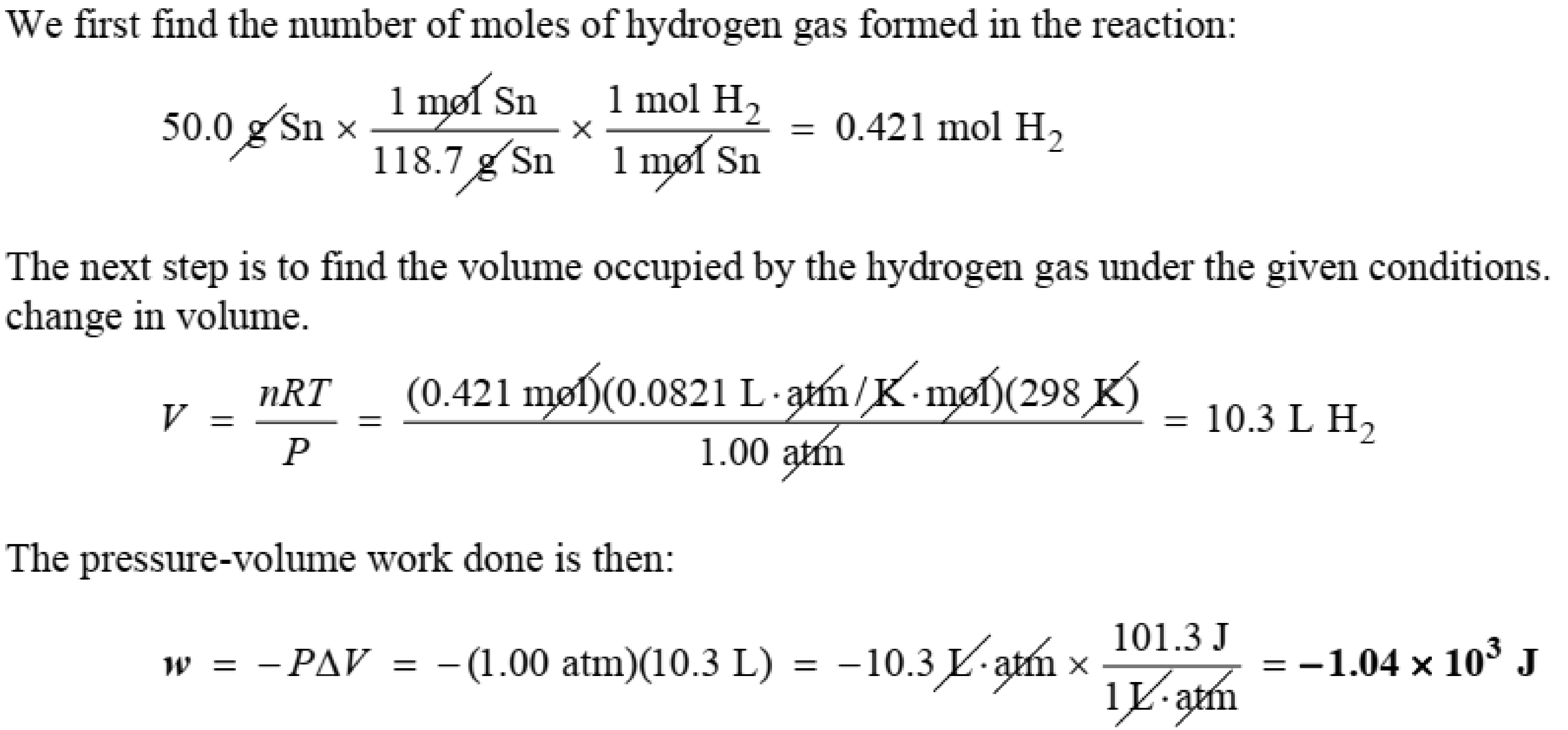
Therefore, we have:

$$\Delta E = q + w = -26\ \text{J} + 74\ \text{J} = 48\ \text{J}$$

6.19 **Calculate the work done when 50.0 g of tin dissolves in excess acid at 1.00 atm and 25°C:**

$Sn(s) + 2H^+(aq)$ ------- $Sn^{2+}(aq) + H_2(g)$

Assume ideal gas behavior.



We first find the number of moles of hydrogen gas formed in the reaction:

$$50.0\ \cancel{g}\ Sn \times \frac{1\ \cancel{mol}\ Sn}{118.7\ \cancel{g}\ Sn} \times \frac{1\ mol\ H_2}{1\ \cancel{mol}\ Sn} = 0.421\ mol\ H_2$$

The next step is to find the volume occupied by the hydrogen gas under the given conditions. This is the change in volume.

$$V = \frac{nRT}{P} = \frac{(0.421\ \cancel{mol})(0.0821\ L\cdot\cancel{atm}/\cancel{K}\cdot\cancel{mol})(298\ \cancel{K})}{1.00\ \cancel{atm}} = 10.3\ L\ H_2$$

The pressure-volume work done is then:

$$w = -P\Delta V = -(1.00\ atm)(10.3\ L) = -10.3\ \cancel{L}\cdot\cancel{atm} \times \frac{101.3\ J}{1\ \cancel{L}\cdot\cancel{atm}} = \mathbf{-1.04\times10^3\ J}$$

6.20 Calculate the work done in joules when 1.0 mole of water vaporizes at 1.0 atm and 100°C. Assume that the volume of liquid water is negligible compared with that of steam at 100°C, and ideal gas behavior

Volume of the water vapor (V_f). volume of liquid water (V_i) is zero
Using the ideal gas equation.

$$V_{H_2O} = \frac{n_{H_2O}RT}{P} = \frac{(1\ \cancel{\text{mol}})\left(0.0821\frac{\text{L}\cdot\cancel{\text{atm}}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(373\ \cancel{\text{K}})}{(1.0\ \cancel{\text{atm}})} = 31\ \text{L}$$

$$\Delta V = V_f - V_i = 31\ \text{L} - 0\ \text{L} = 31\ \text{L}$$

$$w = -P\Delta V = -(1.0\ \text{atm})(31\ \text{L}) = -31\ \text{L}\cdot\text{atm}$$

$$w = -31\ \cancel{\text{L}}\cdot\cancel{\text{atm}} \times \frac{101.3\ \text{J}}{1\ \cancel{\text{L}}\cdot\cancel{\text{atm}}} = \mathbf{-3.1\times10^{3}\ J}$$

6.25 The first step in the industrial recovery of zinc from the zinc sulfide ore is roasting, that is, the conversion of ZnS to ZnO by heating:

$$2\mathrm{ZnS}(s) + 3\mathrm{O}_2(g) \longrightarrow 2\mathrm{ZnO}(s) + 2\mathrm{SO}_2(g) \qquad \Delta H = -879\ \mathrm{kJ/mol}$$

Calculate the heat evolved (in kJ) per gram of ZnS roasted.

$$1\mathrm{g\ ZnS} \times \frac{1\ \text{mole ZnS}}{97.46\ \text{g ZnS}} \times \frac{879\ \text{KJ}}{2\ \text{mole ZnS}} = 4.51\ \mathrm{kJ\ /\ g\ ZnS}$$

6.26 Determine the amount of heat (in kJ) given off when 1.26 x 10^4 g of NO_2 are produced according to the equation

$$2\mathrm{NO}(g) + \mathrm{O}_2(g) \longrightarrow 2\mathrm{NO}_2(g) \qquad \Delta H = -114.6 \text{ kJ/mol}$$

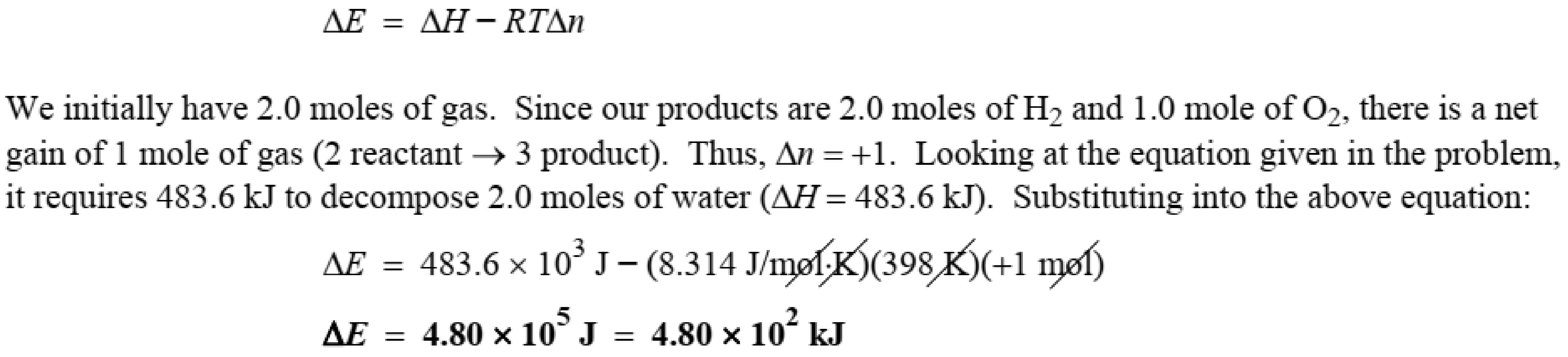
$$(1.26 \times 10^4 \cancel{\text{g}}\ \mathrm{NO}_2) \times \frac{1\ \cancel{\text{mol}}\ \mathrm{NO}_2}{46.01\ \cancel{\text{g}}\ \mathrm{NO}_2} \times \frac{114.6\ \text{kJ}}{2\ \cancel{\text{mol}}\ \mathrm{NO}_2} = \mathbf{1.57 \times 10^4\ kJ}$$

6.27 Consider the reaction

$$2H_2O(g) \longrightarrow 2H_2(g) + O_2(g)$$

$$\Delta H = 483.6 \text{ kJ/mol}$$

If 2.0 moles of $H_2O(g)$ are converted to $H_2(g)$ and $O_2(g)$ against a pressure of 1.0 atm at 125°C, what is ΔE for this reaction?



$$\Delta E = \Delta H - RT\Delta n$$

We initially have 2.0 moles of gas. Since our products are 2.0 moles of H_2 and 1.0 mole of O_2, there is a net gain of 1 mole of gas (2 reactant → 3 product). Thus, $\Delta n = +1$. Looking at the equation given in the problem, it requires 483.6 kJ to decompose 2.0 moles of water ($\Delta H = 483.6$ kJ). Substituting into the above equation:

$$\Delta E = 483.6 \times 10^3 \text{ J} - (8.314 \text{ J/}\cancel{\text{mol}}\cdot\cancel{\text{K}})(398\,\cancel{\text{K}})(+1\ \cancel{\text{mol}})$$

$$\mathbf{\Delta E = 4.80 \times 10^5 \text{ J} = 4.80 \times 10^2 \text{ kJ}}$$

6.32 A piece of silver of mass 362 g has a heat capacity of 85.7 J/°C. What is the specific heat of silver?

$$\text{Specific heat} = \frac{C}{m} = \frac{85.7\ \text{J/°C}}{362\ \text{g}} = \mathbf{0.237\ J/g \cdot °C}$$

6.33 A 6.22-kg piece of copper metal is heated from 20.5°C to 324.3°C. Calculate the heat absorbed (in kJ) by the metal.

$$q = m_{\text{Cu}} s_{\text{Cu}} \Delta t = (6.22 \times 10^3\ \cancel{\text{g}})(0.385\ \text{J}/\cancel{\text{g}} \cdot \cancel{°\text{C}})(324.3 - 20.5)\cancel{°\text{C}} = 7.28 \times 10^5\ \text{J} = \mathbf{728\ kJ}$$

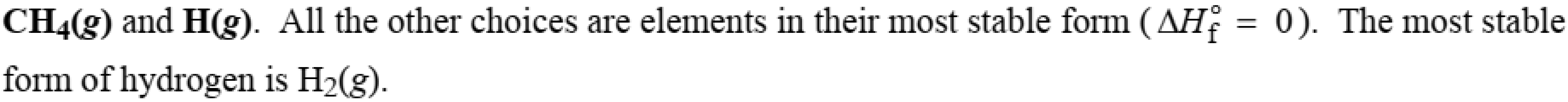
6.34 Calculate the amount of heat liberated (in kJ) from 366 g of mercury when it cools from 77.0°C to 12.0°C.

$$q = ms\Delta t = (366\ \cancel{\text{g}})(0.139\ \text{J}/\cancel{\text{g}}\cdot\cancel{{}^\circ\text{C}})(12.0 - 77.0)\cancel{{}^\circ\text{C}} = -3.31 \times 10^{3}\ \text{J} = -3.31\ \text{kJ}$$

The amount of heat *liberated* is **3.31 kJ**.

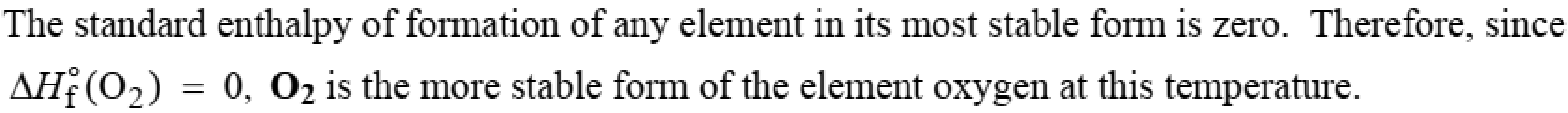
6.45 Which of the following standard enthalpy of formation values is not zero at 25°C?

Na(*s*), Ne(*g*), CH_4(*g*), S_8(*s*), Hg(*l*), H(*g*).



CH_4(*g*) and H(*g*). All the other choices are elements in their most stable form ($\Delta H_f^\circ = 0$). The most stable form of hydrogen is H_2(g).

6.46 The Δ *H*°f values of the two allotropes of oxygen, O_2 and O_3, are 0 and 142.2 kJ/mol, respectively, at 25°C.
Which is the more stable form at this temperature?



The standard enthalpy of formation of any element in its most stable form is zero. Therefore, since $\Delta H_f^\circ(O_2) = 0$, **$O_2$** is the more stable form of the element oxygen at this temperature.

6.47 Which is the more negative quantity at 25°C: Δ *H*°f for $H_2O(l)$ or Δ *H*°f for $H_2O(g)$?

$$H_2O(l) \rightarrow H_2O(g) \qquad \text{Endothermic}$$

$$\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_f[H_2O(g)] - \Delta H^\circ_f[H_2O(l)] > 0$$

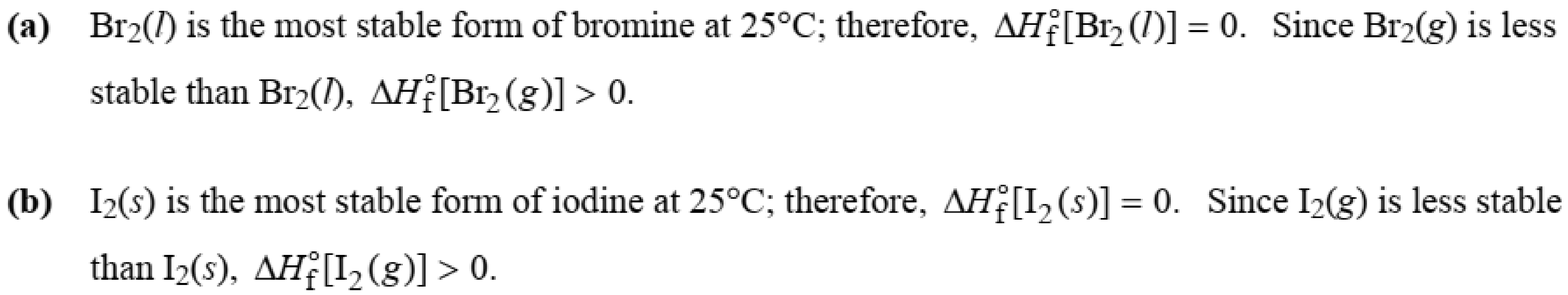
$\Delta H^\circ_f[H_2O(l)]$ is more negative since $\Delta H^\circ_{\text{rxn}} > 0$.

$H_2O(l)$ is the stable form of water at 25°C, and therefore will have the more negative Δ*H*°f value.

6.48 Predict the value of Δ *H*°f (greater than, less than, or equal to zero) for these elements at 25°C

(a) $Br_2(g)$; $Br_2(l)$

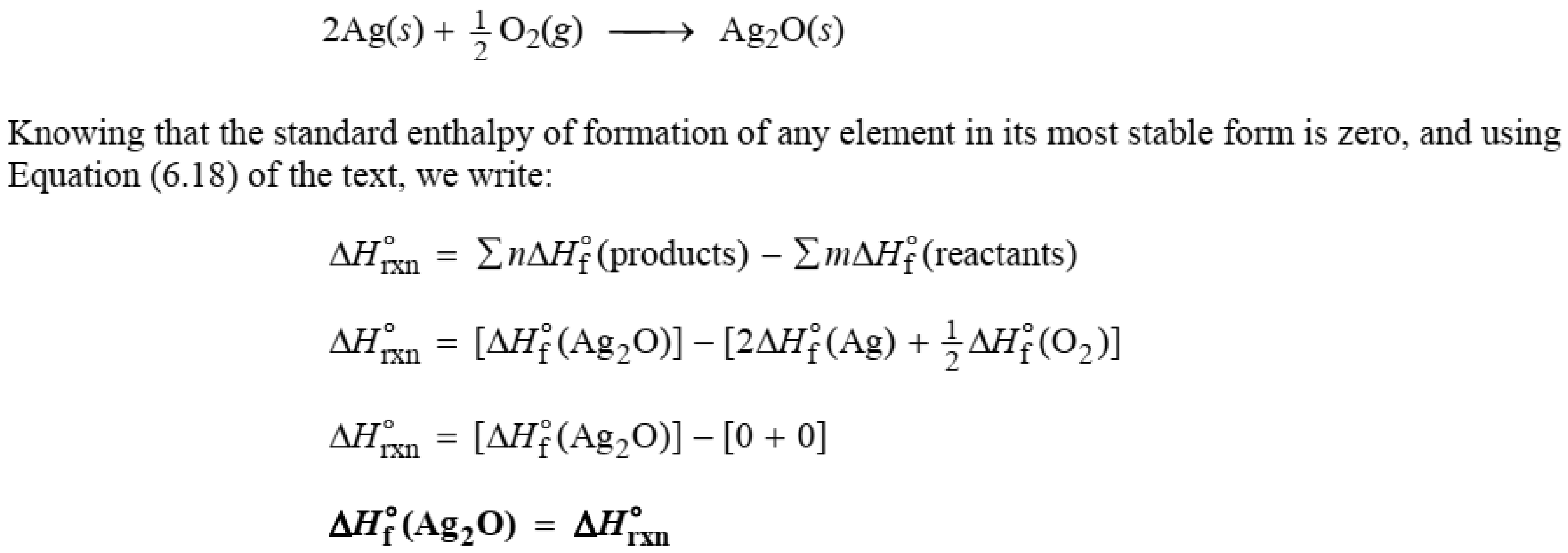
(b) $I_2(g)$; $I_2(s)$.



(a) $Br_2(l)$ is the most stable form of bromine at 25°C; therefore, $\Delta H^\circ_f[Br_2(l)] = 0$. Since $Br_2(g)$ is less stable than $Br_2(l)$, $\Delta H^\circ_f[Br_2(g)] > 0$.

(b) $I_2(s)$ is the most stable form of iodine at 25°C; therefore, $\Delta H^\circ_f[I_2(s)] = 0$. Since $I_2(g)$ is less stable than $I_2(s)$, $\Delta H^\circ_f[I_2(g)] > 0$.

6.50 Suggest ways that would enable you to measure the Δ*H*°f values of $Ag_2O(s)$ from its elements. No calculations are necessary.



$$2Ag(s) + \frac{1}{2}O_2(g) \longrightarrow Ag_2O(s)$$

Knowing that the standard enthalpy of formation of any element in its most stable form is zero, and using Equation (6.18) of the text, we write:

$$\Delta H^\circ_{\text{rxn}} = \sum n\Delta H^\circ_{\text{f}}(\text{products}) - \sum m\Delta H^\circ_{\text{f}}(\text{reactants})$$

$$\Delta H^\circ_{\text{rxn}} = [\Delta H^\circ_{\text{f}}(Ag_2O)] - [2\Delta H^\circ_{\text{f}}(Ag) + \frac{1}{2}\Delta H^\circ_{\text{f}}(O_2)]$$

$$\Delta H^\circ_{\text{rxn}} = [\Delta H^\circ_{\text{f}}(Ag_2O)] - [0 + 0]$$

$$\boldsymbol{\Delta H^\circ_{\text{f}}(Ag_2O) = \Delta H^\circ_{\text{rxn}}}$$

6.51 Calculate the heat of decomposition for this process at constant pressure and 25°C:

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

$$\Delta H° = [\Delta H_f^\circ(CaO) + \Delta H_f^\circ(CO_2)] - \Delta H_f^\circ(CaCO_3)$$

$$\boldsymbol{\Delta H°} = [(1)(-635.6\text{ kJ/mol}) + (1)(-393.5\text{ kJ/mol})] - (1)(-1206.9\text{ kJ/mol}) = \mathbf{177.8\text{ kJ/mol}}$$

6.52 The standard enthalpies of formation of ions in aqueous solutions are obtained by arbitrarily assigning a value of zero to H^+ ions; that is, $\Delta H_f^\circ[H^+(aq)] = 0$.

(a) For the following reaction

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq) \qquad \Delta H^\circ = -74.9\ \text{kJ/mol}$$

calculate ΔH_f° for the Cl^- ions.

(a) $HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$

$$\Delta H_{rxn}^\circ = \Delta H_f^\circ(H^+) + \Delta H_f^\circ(Cl^-) - \Delta H_f^\circ(HCl)$$

$$-74.9\ \text{kJ/mol} = 0 + \Delta H_f^\circ(Cl^-) - (1)(-92.3\ \text{kJ/mol})$$

$$\boldsymbol{\Delta H_f^\circ(Cl^-) = -167.2\ \text{kJ/mol}}$$

6.53 Calculate the heats of combustion for the following reactions from the standard enthalpies of formation listed in Appendix 3:

(a) $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

(b) $2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(l)$

(a) $\Delta H^\circ = 2\Delta H^\circ_f(H_2O) - 2\Delta H^\circ_f(H_2) - \Delta H^\circ_f(O_2)$

$\boldsymbol{\Delta H^\circ} = (2)(-285.8\text{ kJ/mol}) - (2)(0) - (1)(0) = \mathbf{-571.6\text{ kJ/mol}}$

(b) $\Delta H^\circ = 4\Delta H^\circ_f(CO_2) + 2\Delta H^\circ_f(H_2O) - 2\Delta H^\circ_f(C_2H_2) - 5\Delta H^\circ_f(O_2)$

$\boldsymbol{\Delta H^\circ} = (4)(-393.5\text{ kJ/mol}) + (2)(-285.8\text{ kJ/mol}) - (2)(226.6\text{ kJ/mol}) - (5)(0) = \mathbf{-2599\text{ kJ/mol}}$

6.54 Calculate the heats of combustion for the following reactions from the standard enthalpies of formation listed in Appendix 3:

(a) $C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l)$

(b) $2H_2S(g) + 3O_2(g) \longrightarrow 2H_2O(l) + 2SO_2(g)$

(a) $\Delta H^\circ = [2\Delta H_f^\circ(CO_2) + 2\Delta H_f^\circ(H_2O)] - [\Delta H_f^\circ(C_2H_4) + 3\Delta H_f^\circ(O_2)]]$

$\Delta H^\circ = [(2)(-393.5\ \text{kJ/mol}) + (2)(-285.8\ \text{kJ/mol})] - [(1)(52.3\ \text{kJ/mol}) + (3)(0)]$

$\boldsymbol{\Delta H^\circ = -1411\ \textbf{kJ/mol}}$

(b) $\Delta H^\circ = [2\Delta H_f^\circ(H_2O) + 2\Delta H_f^\circ(SO_2)] - [2\Delta H_f^\circ(H_2S) + 3\Delta H_f^\circ(O_2)]$

$\Delta H^\circ = [(2)(-285.8\ \text{kJ/mol}) + (2)(-296.1\ \text{kJ/mol})] - [(2)(-20.15\ \text{kJ/mol}) + (3)(0)]$

$\boldsymbol{\Delta H^\circ = -1124\ \textbf{kJ/mol}}$

6.55 Methanol, ethanol, and *n*-propanol are three common alcohols. When 1.00 g of each of these alcohols is burned in air, heat is liberated as shown by the following data: (a) methanol (CH_3OH), −22.6 kJ; (b) ethanol (C_2H_5OH), −29.7 kJ; (c) *n*-propanol (C_3H_7OH), −33.4 kJ. Calculate the heats of combustion of these alcohols in kJ/mol.

The given enthalpies are in units of kJ/g. We must convert them to units of kJ/mol.

(a) $$\frac{-22.6\text{ kJ}}{1\,\cancel{\text{g}}} \times \frac{32.04\,\cancel{\text{g}}}{1\text{ mol}} = \mathbf{-724\text{ kJ/mol}}$$

(b) $$\frac{-29.7\text{ kJ}}{1\,\cancel{\text{g}}} \times \frac{46.07\,\cancel{\text{g}}}{1\text{ mol}} = \mathbf{-1.37 \times 10^3\text{ kJ/mol}}$$

(c) $$\frac{-33.4\text{ kJ}}{1\,\cancel{\text{g}}} \times \frac{60.09\,\cancel{\text{g}}}{1\text{ mol}} = \mathbf{-2.01 \times 10^3\text{ kJ/mol}}$$

6.56 The standard enthalpy change for the following reaction is 436.4 kJ/mol:

$$H_2(g) \longrightarrow H(g) + H(g)$$

Calculate the standard enthalpy of formation of atomic hydrogen (H).

$$\Delta H^\circ_{\text{rxn}} = \sum n\Delta H^\circ_f(\text{products}) - \sum m\Delta H^\circ_f(\text{reactants})$$

The reaction is:

$$H_2(g) \longrightarrow H(g) + H(g)$$

and,

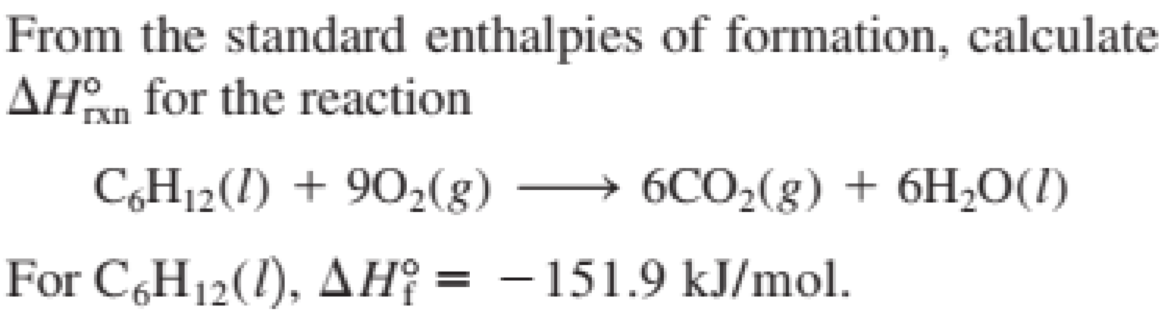
$$\Delta H^\circ_{\text{rxn}} = [\Delta H^\circ_f(\text{H}) + \Delta H^\circ_f(\text{H})] - \Delta H^\circ_f(\text{H}_2)$$

$$\Delta H^\circ_f(\text{H}_2) = 0$$

$$\Delta H^\circ_{\text{rxn}} = 436.4\ \text{kJ/mol} = 2\Delta H^\circ_f(\text{H}) - (1)(0)$$

$$\boldsymbol{\Delta H^\circ_f(\text{H})} = \frac{436.4\ \text{kJ/mol}}{2} = \mathbf{218.2\ kJ/mol}$$

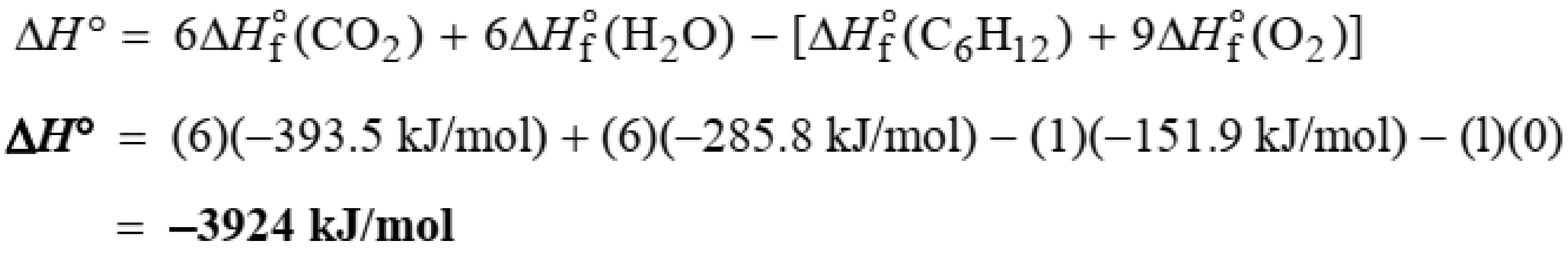
6.57



From the standard enthalpies of formation, calculate $\Delta H^\circ_{\text{rxn}}$ for the reaction

$$C_6H_{12}(l) + 9O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$$

For $C_6H_{12}(l)$, $\Delta H^\circ_f = -151.9$ kJ/mol.

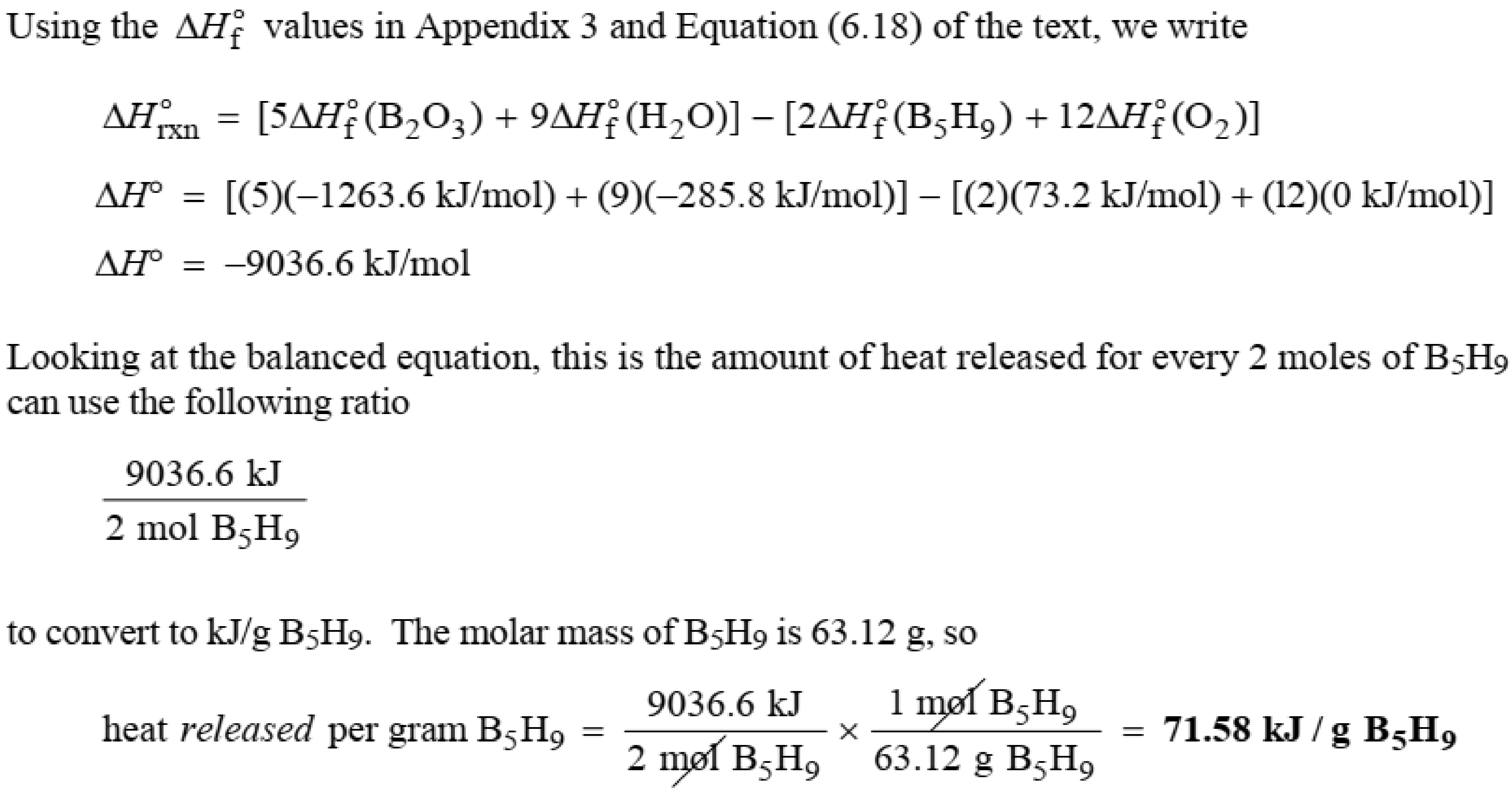


$$\Delta H^\circ = 6\Delta H^\circ_f(CO_2) + 6\Delta H^\circ_f(H_2O) - [\Delta H^\circ_f(C_6H_{12}) + 9\Delta H^\circ_f(O_2)]$$

$$\boldsymbol{\Delta H^\circ} = (6)(-393.5\text{ kJ/mol}) + (6)(-285.8\text{ kJ/mol}) - (1)(-151.9\text{ kJ/mol}) - (1)(0)$$

$$= \mathbf{-3924\ kJ/mol}$$

6.58 Pentaborane-9, B_5H_9, is a colorless, highly reactive liquid that will burst into flame when exposed to oxygen. The reaction is $2B_5H_9(l) + 12O_2(g) \longrightarrow 5B_2O_3(s) + 9H_2O(l)$ **Calculate the kilojoules of heat released per gram of the compound reacted with oxygen. The standard enthalpy of formation of B5H9 is 73.2 kJ/mol.**



Using the ΔH_f° values in Appendix 3 and Equation (6.18) of the text, we write

$$\Delta H_{rxn}^\circ = [5\Delta H_f^\circ(B_2O_3) + 9\Delta H_f^\circ(H_2O)] - [2\Delta H_f^\circ(B_5H_9) + 12\Delta H_f^\circ(O_2)]$$

$$\Delta H^\circ = [(5)(-1263.6\text{ kJ/mol}) + (9)(-285.8\text{ kJ/mol})] - [(2)(73.2\text{ kJ/mol}) + (12)(0\text{ kJ/mol})]$$

$$\Delta H^\circ = -9036.6\text{ kJ/mol}$$

Looking at the balanced equation, this is the amount of heat released for every 2 moles of B_5H_9 reacted. We can use the following ratio

$$\frac{9036.6\text{ kJ}}{2\text{ mol }B_5H_9}$$

to convert to kJ/g B_5H_9. The molar mass of B_5H_9 is 63.12 g, so

$$\text{heat } \textit{released} \text{ per gram } B_5H_9 = \frac{9036.6\text{ kJ}}{2\text{ }\cancel{\text{mol}}\text{ }B_5H_9} \times \frac{1\text{ }\cancel{\text{mol}}\text{ }B_5H_9}{63.12\text{ g }B_5H_9} = \mathbf{71.58\text{ kJ / g }B_5H_9}$$

6.59 Determine the amount of heat (in kJ) given off when 1.26×10^4 g of ammonia are produced according to the equation

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g) \qquad \Delta H^\circ_{\text{rxn}} = -92.6 \text{ kJ/mol}$$

The amount of heat given off is:

$$(1.26 \times 10^4 \cancel{\text{g}}\ \text{NH}_3) \times \frac{1\ \cancel{\text{mol}}\ \text{NH}_3}{17.03\ \cancel{\text{g}}\ \text{NH}_3} \times \frac{92.6\ \text{kJ}}{2\ \cancel{\text{mol}}\ \text{NH}_3} = \mathbf{3.43 \times 10^4\ kJ}$$

6.61 From these data,

$$\mathrm{S(rhombic)} + \mathrm{O_2}(g) \longrightarrow \mathrm{SO_2}(g) \qquad \Delta H^\circ_{\text{rxn}} = -296.06\ \text{kJ/mol}$$

$$\mathrm{S(monoclinic)} + \mathrm{O_2}(g) \longrightarrow \mathrm{SO_2}(g) \qquad \Delta H^\circ_{\text{rxn}} = -296.36\ \text{kJ/mol}$$

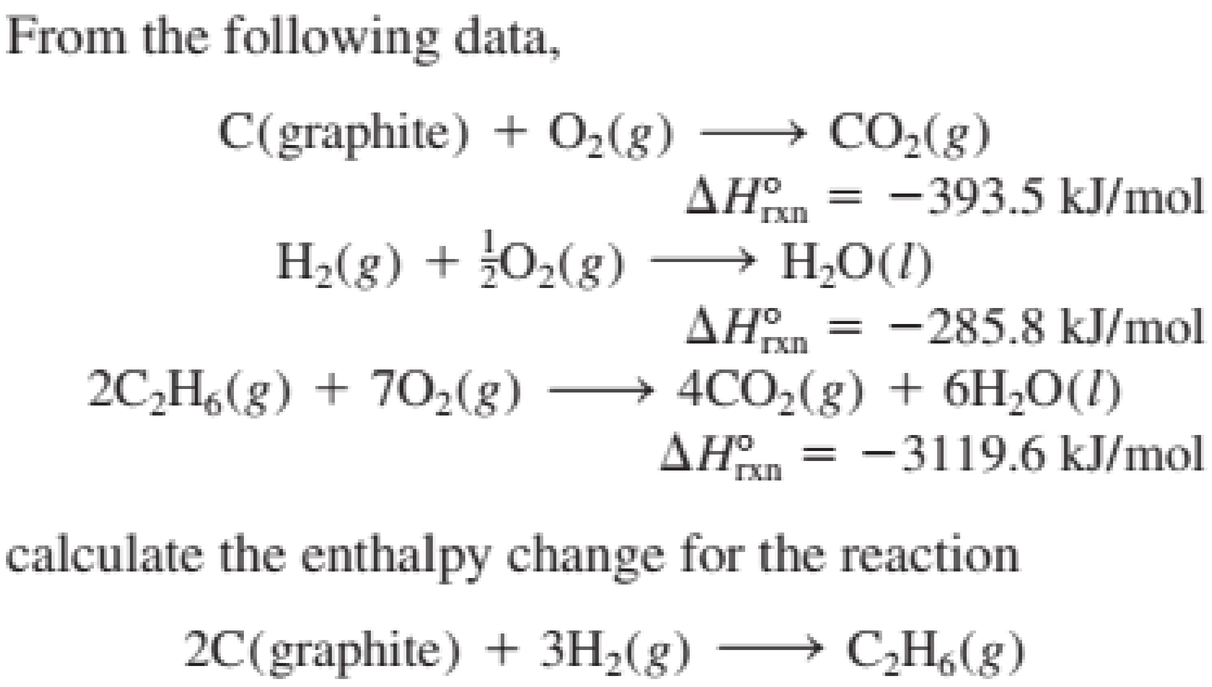
calculate the enthalpy change for the transformation

$$\mathrm{S(rhombic)} \longrightarrow \mathrm{S(monoclinic)}$$

(Monoclinic and rhombic are different allotropic forms of elemental sulfur.)

Reaction	ΔH° (kJ/mol)
$\mathrm{S(rhombic)} + \mathrm{O_2}(g) \rightarrow \mathrm{SO_2}(g)$	–296.06
$\mathrm{SO_2}(g) \rightarrow \mathrm{S(monoclinic)} + \mathrm{O_2}(g)$	296.36
$\mathrm{S(rhombic)} \rightarrow \mathrm{S(monoclinic)}$	$\Delta H^\circ_{\text{rxn}}$ = **0.30 kJ/mol**

6.62



From the following data,

$$C(\text{graphite}) + O_2(g) \longrightarrow CO_2(g)$$
$$\Delta H^\circ_{\text{rxn}} = -393.5 \text{ kJ/mol}$$
$$H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l)$$
$$\Delta H^\circ_{\text{rxn}} = -285.8 \text{ kJ/mol}$$
$$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(l)$$
$$\Delta H^\circ_{\text{rxn}} = -3119.6 \text{ kJ/mol}$$

calculate the enthalpy change for the reaction

$$2C(\text{graphite}) + 3H_2(g) \longrightarrow C_2H_6(g)$$

6.62

$$2C(\text{graphite}) + 3H_2(g) \longrightarrow C_2H_6(g) \qquad \Delta H^\circ_{\text{rxn}} = ?$$

Looking at this reaction, we need two moles of graphite as a reactant. So, we multiply Equation (a) by two to obtain:

$$\text{(d)} \quad 2C(\text{graphite}) + 2O_2(g) \longrightarrow 2CO_2(g) \qquad \Delta H^\circ_{\text{rxn}} = 2(-393.5\text{ kJ/mol}) = -787.0\text{ kJ/mol}$$

Next, we need three moles of H_2 as a reactant. So, we multiply Equation (b) by three to obtain:

$$\text{(e)} \quad 3H_2(g) + \tfrac{3}{2}O_2(g) \longrightarrow 3H_2O(l) \qquad \Delta H^\circ_{\text{rxn}} = 3(-285.8\text{ kJ/mol}) = -857.4\text{ kJ/mol}$$

Last, we need one mole of C_2H_6 as a product. Equation (c) has two moles of C_2H_6 as a reactant, so we need to reverse the equation and divide it by 2.

$$\text{(f)} \quad 2CO_2(g) + 3H_2O(l) \longrightarrow C_2H_6(g) + \tfrac{7}{2}O_2(g) \qquad \Delta H^\circ_{\text{rxn}} = \tfrac{1}{2}(3119.6\text{ kJ/mol}) = 1559.8\text{ kJ/mol}$$

Adding Equations (d), (e), and (f) together, we have:

Reaction	ΔH° (kJ/mol)
(d) $2C(\text{graphite}) + 2\cancel{O_2}(g) \longrightarrow 2\cancel{CO_2}(g)$	−787.0
(e) $3H_2(g) + \tfrac{3}{2}\cancel{O_2}(g) \longrightarrow 3\cancel{H_2O}(l)$	−857.4
(f) $2\cancel{CO_2}(g) + 3\cancel{H_2O}(l) \longrightarrow C_2H_6(g) + \tfrac{7}{2}\cancel{O_2}(g)$	1559.8
$2C(\text{graphite}) + 3H_2(g) \longrightarrow C_2H_6(g)$	**$\Delta H^\circ = -84.6$ kJ/mol**

6.63 From the following heats of combustion,

$$CH_3OH(l) + \tfrac{3}{2}O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H^\circ_{rxn} = -726.4\ \text{kJ/mol}$$

$$C(\text{graphite}) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H^\circ_{rxn} = -393.5\ \text{kJ/mol}$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l) \qquad \Delta H^\circ_{rxn} = -285.8\ \text{kJ/mol}$$

calculate the enthalpy of formation of methanol (CH_3OH) from its elements:

$$C(\text{graphite}) + 2H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow CH_3OH(l)$$

Reaction	ΔH° (kJ/mol)
$\cancel{CO_2}(g) + \cancel{2H_2O}(l) \rightarrow CH_3OH(l) + \frac{3}{2}O_2(g)$	726.4
$C(\text{graphite}) + O_2(g) \rightarrow \cancel{CO_2}(g)$	–393.5
$2H_2(g) + O_2(g) \rightarrow \cancel{2H_2O}(l)$	2(–285.8)
$C(\text{graphite}) + 2H_2(g) + \frac{1}{2}O_2(g) \rightarrow CH_3OH(l)$	$\Delta H^\circ_{rxn} = $ **–238.7 kJ/mol**

We have just calculated an enthalpy at standard conditions, which we abbreviate ΔH°_{rxn}. In this case, the reaction in question was for the formation of *one* mole of CH_3OH *from its elements* in their standard state. Therefore, the ΔH°_{rxn} that we calculated is also, by definition, the standard heat of formation ΔH°_f of CH_3OH (**–238.7 kJ/mol**).

6.64 Calculate the standard enthalpy change for the reaction

$$2\mathrm{Al}(s) + \mathrm{Fe_2O_3}(s) \longrightarrow 2\mathrm{Fe}(s) + \mathrm{Al_2O_3}(s)$$

given that

$$2\mathrm{Al}(s) + \tfrac{3}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{Al_2O_3}(s) \qquad \Delta H^\circ_{\mathrm{rxn}} = -1669.8\ \mathrm{kJ/mol}$$

$$2\mathrm{Fe}(s) + \tfrac{3}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{Fe_2O_3}(s) \qquad \Delta H^\circ_{\mathrm{rxn}} = -822.2\ \mathrm{kJ/mol}$$

The second and third equations can be combined to give the first equation.

$$2\mathrm{Al}(s) + \tfrac{3}{2}\cancel{\mathrm{O_2}}(g) \longrightarrow \mathrm{Al_2O_3}(s) \qquad \Delta H^\circ = -1669.8\ \mathrm{kJ/mol}$$

$$\mathrm{Fe_2O_3}(s) \longrightarrow 2\mathrm{Fe}(s) + \tfrac{3}{2}\cancel{\mathrm{O_2}}(g) \qquad \Delta H^\circ = 822.2\ \mathrm{kJ/mol}$$

$$2\mathrm{Al}(s) + \mathrm{Fe_2O_3}(s) \longrightarrow 2\mathrm{Fe}(s) + \mathrm{Al_2O_3}(s) \qquad \boldsymbol{\Delta H^\circ = -847.6\ \mathrm{kJ/mol}}$$