

## Suggested Problems: Chapter 9

**9.7:** We know that  $q = mS\Delta T$ . Solving for  $q$  gives

$$q = 28.4 \text{ g} \times 2.093 \text{ J/g}^\circ\text{C} \times \{-1.0^\circ\text{C} - (-23.0^\circ\text{C})\} = 1307 \text{ J} \approx 1310 \text{ J}$$

Note that the specific heat for ice is  $2.093 \text{ J/g}^\circ\text{C}$  instead of  $4.184 \text{ J/g}^\circ\text{C}$ , which is the specific heat for liquid water. In terms of calories, the heat required is

$$1307 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J/g}^\circ\text{C}} = 312.5 \text{ cal} \approx 313 \text{ cal}$$

**9.9:** We know that  $q = mS\Delta T$ . Solving for  $\Delta T$  gives

$$\Delta T = \frac{q}{ms} = \frac{14500 \text{ J}}{485 \text{ g} \times (4.184 \text{ J/g}^\circ\text{C})} = 7.15^\circ\text{C}$$

**9.11:** We know that  $q = mS\Delta T$ . Solving for  $S$  gives

$$S = \frac{q}{m\Delta T} = \frac{8460 \text{ J}}{437.2 \text{ J} \times (68.9^\circ\text{C} - 19.3^\circ\text{C})} = 0.390 \text{ J/g}^\circ\text{C}$$

Given the options in Table 1, copper (Cu) is the best choice.

**9.19:** We know that  $q_{\text{spoon}} = -q_{\text{coffee}}$ , or that

$$m_{\text{spoon}}S_{\text{spoon}}\Delta T_{\text{spoon}} = m_{\text{coffee}}S_{\text{coffee}}\Delta T_{\text{coffee}}$$

Substituting in know values

$$45 \text{ g} \times 0.24 \text{ J/g}^\circ\text{C} \times (T_{\text{final}} - 25^\circ\text{C}) = -180 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C} \times (T_{\text{final}} - 95^\circ\text{C})$$

$$10.8 \text{ J/}^\circ\text{C} \times (T_{\text{final}} - 25^\circ\text{C}) = -753.12 \text{ J/}^\circ\text{C} \times (T_{\text{final}} - 95^\circ\text{C})$$

$$10.8 \times T_{\text{final}} - 270 = -753.12 \times T_{\text{final}} + 71546.4$$

$$763.12 \times T_{\text{final}} = 71816.4$$

$$T_{\text{final}} = 94.1^\circ\text{C} \approx 94^\circ\text{C}$$

The temperature of the coffee, therefore, decreases by just  $1^\circ\text{C}$ .

**9.25:** We know that  $q_{\text{solution}} = mS\Delta T$ . Solving for  $q$  gives

$$q_{\text{solution}} = (150.0 \text{ g water} + 3.0 \text{ g CaCl}_2) \times 4.184 \text{ J/g}^\circ\text{C} \times (25.8^\circ\text{C} - 22.4^\circ\text{C}) = 2177 \text{ J} \approx 2200 \text{ J}$$

Note that the total mass of the solution is the mass of the water and the mass of the  $\text{CaCl}_2$  added to the water. Because  $q_{\text{reaction}} = -q_{\text{solution}}$ , we know that the  $q_{\text{reaction}} = -2200 \text{ J}$  and the reaction is exothermic.

**9.27:** We know that  $q_{\text{solution}} = mS\Delta T$ . Solving for  $q$  gives

$$q_{\text{soln}} = (3.15 \text{ g Ba(OH)}_2 \cdot 8 \text{ H}_2\text{O} + 1.52 \text{ g NH}_4\text{SCN} + 100 \text{ g water}) \times 4.20 \text{ J/g}^\circ\text{C} \times (-3.1^\circ\text{C}) = -1362.8 \text{ J}$$

or approximately  $-1400 \text{ J}$  to two significant figures. Note that the total mass of solution is the sum of the mass of  $\text{Ba(OH)}_2 \cdot 8 \text{ H}_2\text{O}$ ,  $\text{NH}_4\text{SCN}$ , and water. Because  $q_{\text{reaction}} = -q_{\text{solution}}$ , we know that the  $q_{\text{reaction}} = 1400 \text{ J}$  or  $1.4 \text{ kJ}$ .

**9.39:** An enthalpy change is defined in terms of the reaction's stoichiometry, which means that for this reaction, it is the heat released when one mole of  $\text{HCl}$  reacts with one mole of  $\text{NaOH}$ . In Example 9.5, the heat released is from the reaction of  $0.0500$  moles of  $\text{HCl}$  and  $0.0500$  moles of  $\text{NaOH}$ .

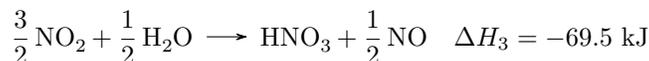
**9.41:** In Example 9.6, the dissolution of  $3.21 \text{ g}$  of solid  $\text{NH}_4\text{NO}_3$  results in a  $q_{\text{reaction}}$  of  $+1.0 \text{ kJ}$ . To convert this to a  $\Delta H$ , we recall that

$$\Delta H = \frac{q_{\text{reaction}}}{\text{mol NH}_4\text{NO}_3} \times \frac{1 \text{ mol NH}_4\text{NO}_3}{\text{mol}_{\text{rxn}}}$$

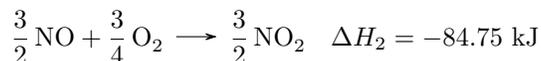
$$\Delta H = \frac{+1.0 \text{ kJ}}{3.21 \text{ g}/(80.043 \text{ g/mol})} \times \frac{1 \text{ mol NH}_4\text{NO}_3}{\text{mol}_{\text{rxn}}} = 25 \text{ kJ/mol}_{\text{rxn}}$$

**9.55:** The standard heat of formation,  $\Delta H_f^\circ$  is the heat released or absorbed when making *one* mole of the substance of interest from its elements in their normal physical state at room temperature and pressure. In this case, the reaction as written is for the formation of *two* moles of  $\text{H}_2\text{O}(g)$ ; thus, the enthalpy change for this reaction is  $2 \times \Delta H_f^\circ$ .

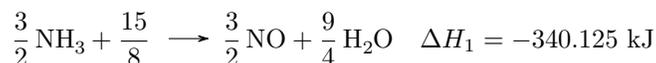
**9.59:** To determine the total energy change per mole of  $\text{HNO}_3$ , we need to adjust the stoichiometry of each reaction to arrive at a net reaction that yields just one mole of  $\text{HNO}_3$ . The last reaction has two moles of  $\text{HNO}_3$  as product, so we divide all coefficients by two and divide the reaction's  $\Delta H$  by two as well



Now, we need the second reaction to produce  $\frac{3}{2} \text{NO}_2$ , which we accomplish by multiplying all coefficients by  $\frac{3}{4}$  and multiplying the reaction's  $\Delta H$  by this value as well



Finally, we need the first reaction to produce  $\frac{3}{2} \text{NO}$ , which we accomplish by multiplying all coefficients by  $\frac{3}{8}$  and multiplying the reaction's  $\Delta H$  by this value as well



The total energy is

$$\Delta H_1 + \Delta H_2 + \Delta H_3 = (-340.125) + (-84.75) + (-69.5) = -494.375 \text{ kJ} \approx -494 \text{ kJ}$$

**9.61:** For this reaction, we know that

$$\Delta H = \Delta H_{f,\text{H}_2\text{O}(g)} - \Delta H_{f,\text{H}_2\text{O}(l)} = -241.82 - (-285.83) = 44.01 \text{ kJ/mol}_{\text{rxn}}$$

As expected, the value is positive as we must add heat, an endothermic process, to evaporate water.

**9.69:** For (a) we have

$$\Delta H = (\Delta H_{f,\text{SiF}_4(g)}^\circ) - (\Delta H_{f,\text{Si}(s)}^\circ + 2 \times \Delta H_{f,\text{F}_2}^\circ) = (-1614.9) - (0 + 2 \times 0) = -1614.9 \text{ kJ/mol}_{\text{rxn}}$$

For (b) we have

$$\Delta H = (\Delta H_{f,\text{CH}_3\text{CO}_2\text{H}(l)}^\circ) - (2 \times \Delta H_{f,\text{C}(s)}^\circ + 2 \times \Delta H_{f,\text{H}_2}^\circ + \Delta H_{f,\text{O}_2(g)}^\circ) = (-485.5) - (2 \times 0 + 2 \times 0 + 0) = -485.5 \text{ kJ/mol}_{\text{rxn}}$$

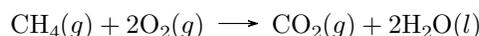
For (c) we have

$$\Delta H = (\Delta H_{f,\text{HCN}(g)}^\circ + \Delta H_{f,\text{NH}_3(g)}^\circ) - (\Delta H_{f,\text{CH}_4(g)}^\circ + \Delta H_{f,\text{N}_2}^\circ) = (135.5 + (-45.9)) - (-74.6 + 0) = +164.2 \text{ kJ/mol}_{\text{rxn}}$$

For (d) we have

$$\Delta H = (\Delta H_{f,\text{CCl}_4(g)}^\circ + \Delta H_{f,\text{S}_2\text{Cl}_2(g)}^\circ) - (\Delta H_{f,\text{CS}_2(g)}^\circ + 3 \times \Delta H_{f,\text{Cl}_2}^\circ) = (-95.7 + (-19.50)) - (116.9 + 3 \times 0) = -232.1 \text{ kJ/mol}_{\text{rxn}}$$

**9.81:** To answer this question we need to (a) balance each reaction, (b) calculate each reaction's  $\Delta H$ , and then (c) convert  $\Delta H$  for each reaction, which is in  $\text{kJ/mol}_{\text{rxn}}$ , to  $\text{kJ/g}$ . The reactions are



The values for  $\Delta H$  are

$$[-1273.5 + (3 \times -241.82)] - [36.4 + (3 \times 0)] = -2035.4 \text{ kJ/mol}_{\text{rxn}}$$

$$[-393.51 + (2 \times -241.82)] - [-74.6 + (2 \times 0)] = -802.6 \text{ kJ/mol}_{\text{rxn}}$$

$$[0 + (2 \times -241.82)] - [50.63 + 0] = -534.27 \text{ kJ/mol}_{\text{rxn}}$$

The molar masses of the three compounds are 27.610 g for  $\text{B}_2\text{H}_6$ , 16.043 g for  $\text{CH}_4$ , and 32.0452 g for  $\text{N}_2\text{H}_4$ . Dividing  $\Delta H$  values by the molar mass gives the heat released per gram of fuel, or  $-70.93 \text{ kJ/g}$  for  $\text{B}_2\text{H}_6$ ,  $-50.03 \text{ kJ/g}$  for  $\text{CH}_4$ , and  $-16.67 \text{ kJ/g}$  for  $\text{N}_2\text{H}_4$ . Of the three fuels,  $\text{B}_2\text{H}_6$  provides the greatest amount of energy per gram.