

$$7 \text{ g protein} \times \frac{17 \text{ kJ}}{1 \text{ g protein}} = 119 \text{ or } 1 \times 10^2 \text{ kJ}$$

$$\text{total energy} = 95.0 \text{ kJ} + 238 \text{ kJ} + 119 \text{ kJ} = 452 \text{ or } 5 \times 10^2 \text{ kJ}$$

$$452 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} \times \frac{1 \text{ Cal}}{1 \text{ kcal}} = 108.03 \text{ or } 1 \times 10^2 \text{ Cal/serving}$$

*Check.* 100 Cal/serving is a reasonable result; units are correct. The data and result have 1 sig fig.

(b) Sodium does not contribute to the calorie content of the food, because it is metabolized by the body; it enters and leaves as  $\text{Na}^+$ .

5.84 Calculate the fuel value in a pound of M&M<sup>®</sup> candies.

$$96 \text{ fat} \times \frac{38 \text{ kJ}}{1 \text{ g fat}} = 3648 \text{ kJ} = 3.6 \times 10^3 \text{ kJ}$$

$$320 \text{ g carbohydrate} \times \frac{17 \text{ kJ}}{1 \text{ g carbohydrate}} = 5440 \text{ kJ} = 5.4 \times 10^3 \text{ kJ}$$

$$21 \text{ g protein} \times \frac{17 \text{ kJ}}{1 \text{ g protein}} = 357 \text{ kJ} = 3.6 \times 10^2 \text{ kJ}$$

$$\text{total fuel value} = 3648 \text{ kJ} + 5440 \text{ kJ} + 357 \text{ kJ} = 9445 \text{ kJ} = 9.4 \times 10^3 \text{ kJ/lb}$$

$$\frac{9445 \text{ kJ}}{\text{lb}} \times \frac{1 \text{ lb}}{453.6 \text{ g}} \times \frac{42 \text{ g}}{\text{serving}} = 874.5 \text{ kJ} = 8.7 \times 10^2 \text{ kJ/serving}$$

$$\frac{874.5 \text{ kJ}}{\text{serving}} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} \times \frac{1 \text{ Cal}}{1 \text{ kcal}} = 209.0 \text{ Cal} = 2.1 \times 10^2 \text{ Cal/serving}$$

*Check.* 210 Cal is the approximate food value of a candy bar, so the result is reasonable.

5.85 *Plan.* g  $\rightarrow$  mol  $\rightarrow$  kJ  $\rightarrow$  Cal *Solve.*

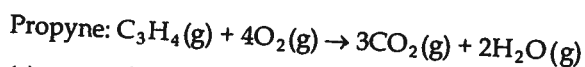
$$16.0 \text{ g C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.2 \text{ g C}_6\text{H}_{12}\text{O}_6} \times \frac{2812 \text{ kJ}}{\text{mol C}_6\text{H}_{12}\text{O}_6} \times \frac{1 \text{ Cal}}{4.184 \text{ kJ}} = 59.7 \text{ Cal}$$

*Check.* 60 Cal is a reasonable result for most of the food value in an apple.

$$5.86 \quad 177 \text{ mL} \times \frac{1.0 \text{ g wine}}{1 \text{ mL}} \times \frac{0.106 \text{ g ethanol}}{1 \text{ g wine}} \times \frac{1 \text{ mol ethanol}}{46.1 \text{ g ethanol}} \times \frac{1367 \text{ kJ}}{1 \text{ mol ethanol}} \times \frac{1 \text{ Cal}}{4.184 \text{ kJ}} \\ = 133 = 1.3 \times 10^2 \text{ Cal}$$

*Check.* A "typical" 6 oz. glass of wine has 150–250 Cal, so this is a reasonable result. Note that alcohol is responsible for most of the food value of wine.

5.87 *Plan.* Use enthalpies of formation to calculate molar heat (enthalpy) of combustion using Hess's Law. Use molar mass to calculate heat of combustion per kg of hydrocarbon. *Solve.*



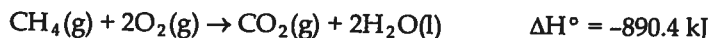
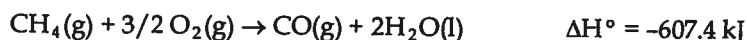
$$(a) \quad \Delta H_{\text{rxn}}^\circ = 3(-393.5 \text{ kJ}) + 2(-241.82 \text{ kJ}) - (185.4 \text{ kJ}) - 4(0) = -1849.5 \\ = -1850 \text{ kJ/mol C}_3\text{H}_4$$

- (b) Use the specific heat of water,  $4.184 \text{ J/g}^\circ\text{C}$ , to calculate the energy required to heat the water. Use the density of water at  $25^\circ\text{C}$  to calculate the mass of  $\text{H}_2\text{O}$  to be heated. (The change in density of  $\text{H}_2\text{O}$  going from  $21^\circ\text{C}$  to  $79^\circ\text{C}$  does not substantially affect the strategy of the exercise.) Then use the 'heat stoichiometry' in (a) to calculate mass of  $\text{Mg(s)}$  needed.

$$75 \text{ mL} \times \frac{0.997 \text{ g H}_2\text{O}}{\text{mL}} \times \frac{4.184 \text{ J}}{\text{g}^\circ\text{C}} \times 58^\circ\text{C} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 18.146 \text{ kJ} = 18 \text{ kJ required}$$

$$18.146 \text{ kJ} \times \frac{1 \text{ mol Mg}}{353.04 \text{ kJ}} \times \frac{24.305 \text{ g Mg}}{1 \text{ mol Mg}} = 1.249 \text{ g} = 1.2 \text{ g Mg needed}$$

- 5.103 (a) For comparison, balance the equations so that 1 mole of  $\text{CH}_4$  is burned in each.



(b)  $\Delta\text{H}_{\text{rxn}}^\circ = \Delta\text{H}_f^\circ \text{C}(\text{s}) + 2\Delta\text{H}_f^\circ \text{H}_2\text{O}(\text{l}) - \Delta\text{H}_f^\circ \text{CH}_4(\text{g}) - \Delta\text{H}_f^\circ \text{O}_2(\text{g})$

$$= 0 + 2(-285.83 \text{ kJ}) - (-74.8) - 0 = -496.9 \text{ kJ}$$

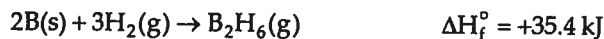
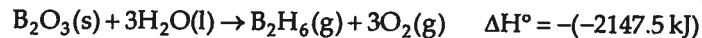
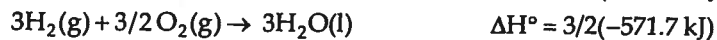
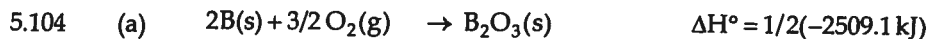
$$\Delta\text{H}_{\text{rxn}}^\circ = \Delta\text{H}_f^\circ \text{CO}(\text{g}) + 2\Delta\text{H}_f^\circ \text{H}_2\text{O}(\text{l}) - \Delta\text{H}_f^\circ \text{CH}_4(\text{g}) - 3/2 \Delta\text{H}_f^\circ \text{O}_2(\text{g})$$

$$= (-110.5 \text{ kJ}) + 2(-285.83 \text{ kJ}) - (-74.8 \text{ kJ}) - 3/2(0) = -607.4 \text{ kJ}$$

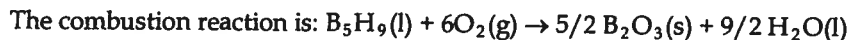
$$\Delta\text{H}_{\text{rxn}}^\circ = \Delta\text{H}_f^\circ \text{CO}_2(\text{g}) + 2\Delta\text{H}_f^\circ \text{H}_2\text{O}(\text{l}) - \Delta\text{H}_f^\circ \text{CH}_4(\text{g}) - 2\Delta\text{H}_f^\circ \text{O}_2(\text{g})$$

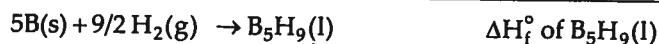
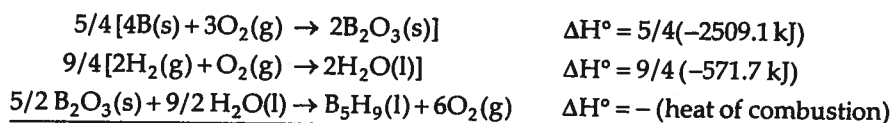
$$= -393.5 \text{ kJ} + 2(-285.83 \text{ kJ}) - (-74.8 \text{ kJ}) - 2(0) = -890.4 \text{ kJ}$$

- (c) Assuming that  $\text{O}_2(\text{g})$  is present in excess, the reaction that produces  $\text{CO}_2(\text{g})$  represents the most negative  $\Delta\text{H}$  per mole of  $\text{CH}_4$  burned. More of the potential energy of the reactants is released as heat during the reaction to give products of lower potential energy. The reaction that produces  $\text{CO}_2(\text{g})$  is the most "downhill" in enthalpy.



- (b) If, like  $\text{B}_2\text{H}_6$ , the combustion of  $\text{B}_5\text{H}_9$  produces  $\text{B}_2\text{O}_3$  as the boron-containing product, the heat of combustion of  $\text{B}_5\text{H}_9$  in addition to data given in part (a) would enable calculation of the heat of formation of  $\text{B}_5\text{H}_9$ .





$$\Delta H_f^\circ \text{ B}_5\text{H}_9(\text{l}) = -[\text{heat of combustion of B}_5\text{H}_9(\text{l})] - 3136.4 \text{ kJ} - 1286 \text{ kJ}$$

We need to measure the heat of combustion of  $\text{B}_5\text{H}_9(\text{l})$ .

5.105 For nitroethane:

$$\frac{1368 \text{ kJ}}{1 \text{ mol C}_2\text{H}_5\text{NO}_2} \times \frac{1 \text{ mol C}_2\text{H}_5\text{NO}_2}{75.072 \text{ g C}_2\text{H}_5\text{NO}_2} \times \frac{1.052 \text{ g C}_2\text{H}_5\text{NO}_2}{1 \text{ cm}^3} = 19.17 \text{ kJ/cm}^3$$

For ethanol:

$$\frac{1367 \text{ kJ}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.069 \text{ g C}_2\text{H}_5\text{OH}} \times \frac{0.789 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ cm}^3} = 23.4 \text{ kJ/cm}^3$$

For methylhydrazine:

$$\frac{1307 \text{ kJ}}{1 \text{ mol CH}_6\text{N}_2} \times \frac{1 \text{ mol CH}_6\text{N}_2}{46.072 \text{ g CH}_6\text{N}_2} \times \frac{0.874 \text{ g CH}_6\text{N}_2}{1 \text{ cm}^3} = 24.8 \text{ kJ/cm}^3$$

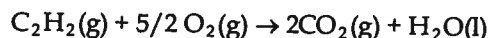
Thus, **methylhydrazine** would provide the most energy per unit volume, with ethanol a close second.

5.106 (a)  $3\text{C}_2\text{H}_2(\text{g}) \rightarrow \text{C}_6\text{H}_6(\text{l})$

$$\Delta H_{\text{rxn}}^\circ = \Delta H_f^\circ \text{C}_6\text{H}_6(\text{l}) - 3\Delta H_f^\circ \text{C}_2\text{H}_2(\text{g}) = 49.0 \text{ kJ} - 3(226.77 \text{ kJ}) = -631.31 = -631.3 \text{ kJ}$$

(b) Since the reaction is exothermic ( $\Delta H$  is negative), the reactant, 3 moles of  $\text{C}_2\text{H}_2(\text{g})$ , has more enthalpy than the product, 1 mole of  $\text{C}_6\text{H}_6(\text{l})$ .

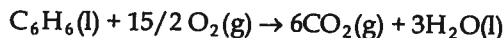
(c) The fuel value of a substance is the amount of heat (kJ) produced when 1 gram of the substance is burned. Calculate the molar heat of combustion (kJ/mol) and use this to find kJ/g of fuel.



$$\Delta H_{\text{rxn}}^\circ = 2\Delta H_f^\circ \text{CO}_2(\text{g}) + \Delta H_f^\circ \text{H}_2\text{O(l)} - \Delta H_f^\circ \text{C}_2\text{H}_2(\text{g}) - 5/2 \Delta H_f^\circ \text{O}_2(\text{g})$$

$$= 2(-393.5 \text{ kJ}) + (-285.83 \text{ kJ}) - 226.77 \text{ kJ} - 5/2(0) = -1299.6 \text{ kJ/mol C}_2\text{H}_2$$

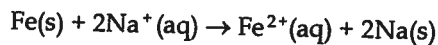
$$\frac{-1299.6 \text{ kJ}}{1 \text{ mol C}_2\text{H}_2} \times \frac{1 \text{ mol C}_2\text{H}_2}{26.036 \text{ g C}_2\text{H}_2} = 49.916 = 50 \text{ kJ/g C}_2\text{H}_2$$



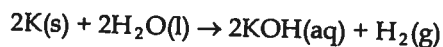
$$\Delta H_{\text{rxn}}^\circ = 6\Delta H_f^\circ \text{CO}_2(\text{g}) + 3\Delta H_f^\circ \text{H}_2\text{O(l)} - \Delta H_f^\circ \text{C}_6\text{H}_6(\text{l}) - 15/2 \Delta H_f^\circ \text{O}_2(\text{g})$$

$$= 6(-393.5 \text{ kJ}) + 3(-285.83 \text{ kJ}) - 49.0 \text{ kJ} - 15/2(0) = -3267.5 \text{ kJ/mol C}_6\text{H}_6$$

$$\frac{-3267.5 \text{ kJ}}{1 \text{ mol C}_6\text{H}_6} \times \frac{1 \text{ mol C}_6\text{H}_6}{78.114 \text{ g C}_6\text{H}_6} = 41.830 = 42 \text{ kJ/g C}_6\text{H}_6$$



$$\begin{aligned}\Delta H^\circ &= \Delta H_f^\circ \text{Fe}^{2+}(\text{aq}) - 2\Delta H_f^\circ \text{Na}^+(\text{aq}) \\ &= -87.86 \text{ kJ} - 2(-240.1 \text{ kJ}) = +392.3 \text{ kJ}\end{aligned}$$



$$\begin{aligned}\Delta H^\circ &= 2\Delta H_f^\circ \text{KOH(aq)} - 2\Delta H_f^\circ \text{H}_2\text{O(l)} \\ &= 2(-482.4 \text{ kJ}) - 2(-285.83 \text{ kJ}) = -393.1 \text{ kJ}\end{aligned}$$

- (c) Exothermic reactions are more likely to be favored, so we expect the first and third reactions be favored.

- (d) In the activity series of metals, Table 4.5, any metal can be oxidized by the cation of a metal below it on the table.

$\text{Ag}^+$  is below Li, so the first reaction will occur.

$\text{Na}^+$  is above Fe, so the second reaction will not occur.

$\text{H}^+$  (formally in  $\text{H}_2\text{O}$ ) is below K, so the third reaction will occur.

These predictions agree with those in part (c).

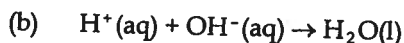
5.114 (a)  $\Delta H^\circ = \Delta H_f^\circ \text{NaNO}_3(\text{aq}) + \Delta H_f^\circ \text{H}_2\text{O(l)} - \Delta H_f^\circ \text{HNO}_3(\text{aq}) - \Delta H_f^\circ \text{NaOH(aq)}$

$$\Delta H^\circ = -446.2 \text{ kJ} - 285.83 \text{ kJ} - (-206.6 \text{ kJ}) - (-469.6 \text{ kJ}) = -55.8 \text{ kJ}$$

$$\Delta H^\circ = \Delta H_f^\circ \text{NaCl(aq)} + \Delta H_f^\circ \text{H}_2\text{O(l)} - \Delta H_f^\circ \text{HCl(aq)} - \Delta H_f^\circ \text{NaOH(aq)}$$

$$\Delta H^\circ = -407.1 \text{ kJ} - 285.83 \text{ kJ} - (-167.2 \text{ kJ}) - (-469.6 \text{ kJ}) = -56.1 \text{ kJ}$$

$$\begin{aligned}\Delta H^\circ &= \Delta H_f^\circ \text{NH}_3(\text{aq}) + \Delta H_f^\circ \text{Na}^+(\text{aq}) + \Delta H_f^\circ \text{H}_2\text{O(l)} - \Delta H_f^\circ \text{NH}_4^+(\text{aq}) - \Delta H_f^\circ \text{NaOH(aq)} \\ &= -80.29 \text{ kJ} - 240.1 \text{ kJ} - 285.83 \text{ kJ} - (-132.5 \text{ kJ}) - (-469.6 \text{ kJ}) = -4.1 \text{ kJ}\end{aligned}$$



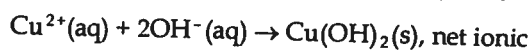
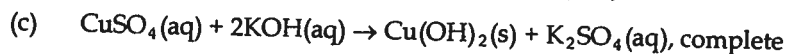
- (c) The  $\Delta H^\circ$  values for the first two reactions are nearly identical,  $-55.8 \text{ kJ}$  and  $-56.1 \text{ kJ}$ . The spectator ions by definition do not change during the course of a reaction, so  $\Delta H^\circ$  is the enthalpy change for the net ionic equation. Since the first two reactions have the same net ionic equation, it is not surprising that they have the same  $\Delta H^\circ$ .

- (d) Strong acids are more likely than weak acids to donate  $\text{H}^+$ . The neutralizations of the two strong acids are energetically favorable, while the neutralization of  $\text{NH}_4^+(\text{aq})$  is significantly less favorable.  $\text{NH}_4^+(\text{aq})$  is probably a weak acid.

5.115 (a)  $\text{mol Cu} = M \times L = 1.00 \text{ M} \times 0.0500 \text{ L} = 0.0500 \text{ mol}$

$$g = \text{mol} \times \text{MM} = 0.0500 \times 63.546 = 3.1773 = 3.18 \text{ g Cu}$$

- (b) The precipitate is copper(II) hydroxide,  $\text{Cu(OH)}_2$ .



- (d) The temperature of the calorimeter rises, so the reaction is exothermic and the sign of  $q$  is negative.

$$q = -6.2^\circ\text{C} \times 100\text{ g} \times \frac{4.184\text{ J}}{1\text{ g}\cdot^\circ\text{C}} = -2.6 \times 10^3\text{ J} = -2.6\text{ kJ}$$

The reaction as carried out involves only 0.050 mol of  $\text{CuSO}_4$  and the stoichiometrically equivalent amount of  $\text{KOH}$ . On a molar basis,

$$\Delta H = \frac{-2.6\text{ kJ}}{0.050\text{ mol}} = -52\text{ kJ for the reaction as written in part (c)}$$



net ionic equation:  $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$

$$\Delta H^\circ = \Delta H_f^\circ \text{AgCl}(\text{s}) - \Delta H_f^\circ \text{Ag}^+(\text{aq}) - \Delta H_f^\circ \text{Cl}^-(\text{aq})$$

$$\Delta H^\circ = -127.0\text{ kJ} - (105.90\text{ kJ}) - (-167.2\text{ kJ}) = -65.7\text{ kJ}$$

- (b)  $\Delta H^\circ$  for the complete molecular equation will be the same as  $\Delta H^\circ$  for the net ionic equation.  $\text{Na}^+(\text{aq})$  and  $\text{NO}_3^-(\text{aq})$  are spectator ions; they appear on both sides of the chemical equation. Since the overall enthalpy change is the enthalpy of the products minus the enthalpy of the reactants, the contributions of the spectator ions cancel.

- (c)  $\Delta H^\circ = \Delta H_f^\circ \text{NaNO}_3(\text{aq}) + \Delta H_f^\circ \text{AgCl}(\text{s}) - \Delta H_f^\circ \text{AgNO}_3(\text{aq}) - \Delta H_f^\circ \text{NaCl}(\text{aq})$   
 $\Delta H_f^\circ \text{AgNO}_3(\text{aq}) = \Delta H_f^\circ \text{NaNO}_3(\text{aq}) + \Delta H_f^\circ \text{AgCl}(\text{s}) - \Delta H_f^\circ \text{NaCl}(\text{aq}) - \Delta H^\circ$   
 $\Delta H_f^\circ \text{AgNO}_3(\text{aq}) = -446.2\text{ kJ} + (-127.0\text{ kJ}) - (-407.1\text{ kJ}) - (-65.7\text{ kJ})$   
 $\Delta H_f^\circ \text{AgNO}_3(\text{aq}) = -100.4\text{ kJ/mol}$

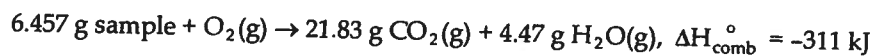
5.117 (a)  $21.83\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \times \frac{1\text{ mol C}}{1\text{ mol CO}_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 5.9572 = 5.957\text{ g C}$   
 $4.47\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}} \times \frac{2\text{ mol H}}{1\text{ mol H}_2\text{O}} \times \frac{1.008\text{ g H}}{\text{mol H}} = 0.5001 = 0.500\text{ g H}$

The sample mass is  $(5.9572 + 0.5001) = 6.457\text{ g}$

(b)  $5.957\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 0.4960\text{ mol C}; 0.4960/0.496 = 1$   
 $0.500\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 0.496\text{ mol H}; 0.496/0.496 = 1$

The empirical formula of the hydrocarbon is  $\text{CH}$ .

- (c) Calculate " $\Delta H_f^\circ$ " for 6.457 g of the sample.



$$\Delta H_{\text{comb}}^\circ = \Delta H_f^\circ \text{CO}_2(\text{g}) + \Delta H_f^\circ \text{H}_2\text{O}(\text{g}) - \Delta H_f^\circ \text{sample} - \Delta H_f^\circ \text{O}_2(\text{g})$$

$$\Delta H_f^\circ \text{sample} = \Delta H_f^\circ \text{CO}_2(\text{g}) + \Delta H_f^\circ \text{H}_2\text{O}(\text{g}) - \Delta H_{\text{comb}}^\circ$$

$$\Delta H_f^\circ \text{CO}_2(\text{g}) = 21.83\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \times \frac{-393.5\text{ kJ}}{\text{mol CO}_2} = -195.185 = -195.2\text{ kJ}$$

- (d) The temperature of the calorimeter rises, so the reaction is exothermic and the sign of  $q$  is negative.

$$q = -6.2^\circ\text{C} \times 100\text{ g} \times \frac{4.184\text{ J}}{1\text{ g}\cdot^\circ\text{C}} = -2.6 \times 10^3\text{ J} = -2.6\text{ kJ}$$

The reaction as carried out involves only 0.050 mol of  $\text{CuSO}_4$  and the stoichiometrically equivalent amount of  $\text{KOH}$ . On a molar basis,

$$\Delta H = \frac{-2.6\text{ kJ}}{0.050\text{ mol}} = -52\text{ kJ for the reaction as written in part (c)}$$



net ionic equation:  $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$

$$\Delta H^\circ = \Delta H_f^\circ \text{AgCl}(\text{s}) - \Delta H_f^\circ \text{Ag}^+(\text{aq}) - \Delta H_f^\circ \text{Cl}^-(\text{aq})$$

$$\Delta H^\circ = -127.0\text{ kJ} - (105.90\text{ kJ}) - (-167.2\text{ kJ}) = -65.7\text{ kJ}$$

- (b)  $\Delta H^\circ$  for the complete molecular equation will be the same as  $\Delta H^\circ$  for the net ionic equation.  $\text{Na}^+(\text{aq})$  and  $\text{NO}_3^-(\text{aq})$  are spectator ions; they appear on both sides of the chemical equation. Since the overall enthalpy change is the enthalpy of the products minus the enthalpy of the reactants, the contributions of the spectator ions cancel.

- (c)  $\Delta H^\circ = \Delta H_f^\circ \text{NaNO}_3(\text{aq}) + \Delta H_f^\circ \text{AgCl}(\text{s}) - \Delta H_f^\circ \text{AgNO}_3(\text{aq}) - \Delta H_f^\circ \text{NaCl}(\text{aq})$   
 $\Delta H_f^\circ \text{AgNO}_3(\text{aq}) = \Delta H_f^\circ \text{NaNO}_3(\text{aq}) + \Delta H_f^\circ \text{AgCl}(\text{s}) - \Delta H_f^\circ \text{NaCl}(\text{aq}) - \Delta H^\circ$   
 $\Delta H_f^\circ \text{AgNO}_3(\text{aq}) = -446.2\text{ kJ} + (-127.0\text{ kJ}) - (-407.1\text{ kJ}) - (-65.7\text{ kJ})$   
 $\Delta H_f^\circ \text{AgNO}_3(\text{aq}) = -100.4\text{ kJ/mol}$

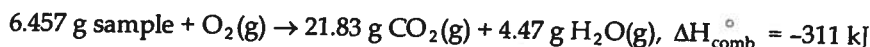
5.117 (a)  $21.83\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \times \frac{1\text{ mol C}}{1\text{ mol CO}_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 5.9572 = 5.957\text{ g C}$   
 $4.47\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}} \times \frac{2\text{ mol H}}{1\text{ mol H}_2\text{O}} \times \frac{1.008\text{ g H}}{\text{mol H}} = 0.5001 = 0.500\text{ g H}$

The sample mass is  $(5.9572 + 0.5001) = 6.457\text{ g}$

(b)  $5.957\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 0.4960\text{ mol C}; 0.4960/0.496 = 1$   
 $0.500\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 0.496\text{ mol H}; 0.496/0.496 = 1$

The empirical formula of the hydrocarbon is  $\text{CH}$ .

- (c) Calculate " $\Delta H_f^\circ$ " for 6.457 g of the sample.



$$\Delta H_{\text{comb}}^\circ = \Delta H_f^\circ \text{CO}_2(\text{g}) + \Delta H_f^\circ \text{H}_2\text{O}(\text{g}) - \Delta H_f^\circ \text{sample} - \Delta H_f^\circ \text{O}_2(\text{g})$$

$$\Delta H_f^\circ \text{sample} = \Delta H_f^\circ \text{CO}_2(\text{g}) + \Delta H_f^\circ \text{H}_2\text{O}(\text{g}) - \Delta H_{\text{comb}}^\circ$$

$$\Delta H_f^\circ \text{CO}_2(\text{g}) = 21.83\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \times \frac{-393.5\text{ kJ}}{\text{mol CO}_2} = -195.185 = -195.2\text{ kJ}$$

$$\Delta H_f^\circ \text{H}_2\text{O}(\text{g}) = 4.47 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{-241.82 \text{ kJ}}{\text{mol H}_2\text{O}} = -59.985 = -60.0 \text{ kJ}$$

$$\Delta H_f^\circ \text{sample} = -195.185 \text{ kJ} - 59.985 \text{ kJ} - (-311 \text{ kJ}) = 55.83 = 56 \text{ kJ}$$

$$H_f^\circ = \frac{55.83 \text{ kJ}}{6.457 \text{ g sample}} \times \frac{13.02 \text{ g}}{\text{CH unit}} = 112.6 = 1.1 \times 10^2 \text{ kJ/CH unit}$$

- (d) The hydrocarbons in Appendix C with empirical formula CH are  $\text{C}_2\text{H}_2$  and  $\text{C}_6\text{H}_6$ .

substance	$\Delta H_f^\circ/\text{mol}$	$\Delta H_f^\circ/\text{CH unit}$
$\text{C}_2\text{H}_2(\text{g})$	226.7 kJ	113.4 kJ
$\text{C}_6\text{H}_6(\text{g})$	82.9 kJ	13.8 kJ
$\text{C}_6\text{H}_6(\text{l})$	49.0 kJ	8.17 kJ
sample		$1.1 \times 10^2 \text{ kJ}$

The calculated value of  $\Delta H_f^\circ/\text{CH unit}$  for the sample is a good match with acetylene,  $\text{C}_2\text{H}_2(\text{g})$ .

- 5.118 (a)  $\text{CH}_4(\text{g}) \rightarrow \text{C}(\text{g}) + 4\text{H}(\text{g})$  (i) reaction given  
 $\text{CH}_4(\text{g}) \rightarrow \text{C}(\text{s}) + 2\text{H}_2(\text{g})$  (ii) reverse of formation

The differences are: the state of C in the products; the chemical form, atoms, or diatomic molecules, of H in the products.

- (b) i.  $\Delta H^\circ = \Delta H_f^\circ \text{C}(\text{g}) + 4\Delta H_f^\circ \text{H}(\text{g}) - \Delta H_f^\circ \text{CH}_4(\text{g})$   
 $= 718.4 \text{ kJ} + 4(217.94) \text{ kJ} - (-74.8) \text{ kJ} = 1665.0 \text{ kJ}$   
 ii.  $\Delta H^\circ = \Delta H_f^\circ \text{CH}_4 = -(-74.8) \text{ kJ} = 74.8 \text{ kJ}$

The rather large difference in  $\Delta H^\circ$  values is due to the enthalpy difference between isolated gaseous C atoms and the orderly, bonded array of C atoms in graphite,  $\text{C}(\text{s})$ , as well as the enthalpy difference between isolated H atoms and  $\text{H}_2$  molecules. In other words, it is due to the difference in the enthalpy stored in chemical bonds in  $\text{C}(\text{s})$  and  $\text{H}_2(\text{g})$  versus the corresponding isolated atoms.

- (c)  $\text{CH}_4(\text{g}) + 4\text{F}_2(\text{g}) \rightarrow \text{CF}_4(\text{g}) + 4\text{HF}(\text{g})$   $\Delta H^\circ = -1679.5 \text{ kJ}$

The  $\Delta H^\circ$  value for this reaction was calculated in Solution 5.88.

$$3.45 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times 0.21509 = 0.215 \text{ mol CH}_4$$

$$1.22 \text{ g F}_2 \times \frac{1 \text{ mol F}_2}{38.00 \text{ g F}_2} = 0.03211 = 0.0321 \text{ mol F}_2$$

There are fewer mol  $\text{F}_2$  than  $\text{CH}_4$ , but 4 mol  $\text{F}_2$  are required for every 1 mol of  $\text{CH}_4$  reacted, so clearly  $\text{F}_2$  is the limiting reactant.