## Chapter 5 Review, pages 338–339

**Knowledge**

**1.** (d)

**2. (**d)

**3.** (a)

**4. (**d)

**5. (**b)

**6.** (c)

**7.** (a)

**8.** (a)

**9.** False. The law of conservation of energy states that the energy of a system is always *constant*.

**10.** True

**11.** True

**12.** False. The molar enthalpy change, $\Delta H_r$, of a substance is the energy change that occurs when *1 mol* of that substance undergoes a physical, chemical, or nuclear change.

**13.** True

**14.** False. *Less* energy is released when a double bond forms than when a triple bond forms.

**15.** False. Bond energies will give *an estimated value* for the enthalpy of a reaction.

**16.** False. Hess's Law states that enthalpy change is *independent* on the number of steps by which a reaction occurs.

**17.** False. If the forward reaction of a reversible chemical reaction has a $\Delta H$ of –62 kJ, the $\Delta H$ of the reverse reaction is *+62 kJ*.

**18.** True

**19.** False. Nuclear generating stations use *fission* to release energy.

**20.** False. Pure ethanol *cannot* be used directly as a fuel source in most automobile engines.

**Understanding**

**21. (a)** The freezing of water is an exothermic reaction.

**(b)** The combustion of acetylene gas is an exothermic reaction.

**(c)** The neutralization of a strong acid with a strong base, which increases the temperature of the solution from 20.5 °C to 26.0 °C is an exothermic reaction.

**(d)** The dissolution of a salt, which changes the temperature of the solution from 21 °C to 19 °C is an endothermic reaction.

**22. (a)** The energy absorbed by the surroundings is 54 kJ.

**(b)** The reaction is exothermic.

**23.** The cocoa in the Thermos stays warm for hours because, unlike the ceramic cup, the Thermos is sealed and insulated like a calorimeter, which has an insulated chamber that minimizes energy loss to the surroundings.

**24.** Methane, gasoline, and propane are effective fuel sources because they are all hydrocarbons in which the molecules contain only C–H and C–C bonds. These are high energy bonds that will break and release a large amount of energy upon combustion.

**25.** We are able to rearrange two chemical equations and add their enthalpy change values to determine the enthalpy change of a third reaction because of Hess's law. Hess's law states that the enthalpy change of the conversion of reactants to products is the same whether the conversion occurs in one step or a series of steps.

**26.** The reaction is exothermic since it releases energy. Therefore, to reverse the reaction, energy must be added. Since we are interested in reversing the reaction using only one-half of the product, we only need one-half the quantity of energy released.
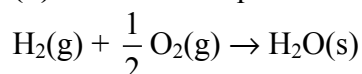
$$\frac{1}{2}(724 \text{ kJ}) = 362 \text{ kJ}$$

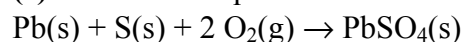Therefore, the reverse reaction will proceed with only the addition of 362 kJ of thermal energy.

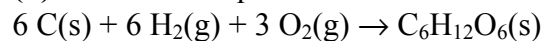**27. (a)** A chemical equation for the formation of NaCl(s) is:

$$\text{Na(s)} + \frac{1}{2}\text{Cl}_2(g) \rightarrow \text{NaCl(s)}$$
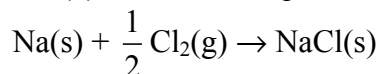
**(b)** A chemical equation for the formation of $H_2O(s)$ is:

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(s)$$

**(c)** A chemical equation for the formation of $PbSO_4(s)$ is:

$Pb(s) + S(s) + 2 O_2(g) \rightarrow PbSO_4(s)$

**(d)** A chemical equation for the formation of $C_6H_{12}O_6(s)$ is:

$6 C(s) + 6 H_2(g) + 3 O_2(g) \rightarrow C_6H_{12}O_6(s)$

**28.** A thermochemical equation for the formation of aluminum oxide, $Al_2O_3(s)$ with the energy term separate is:

$$2 \text{ Al(s)} + \frac{3}{2}O_2(g) \rightarrow Al_2O_3(s) \quad \Delta H_f^\circ = -1676 \text{ kJ/mol}$$
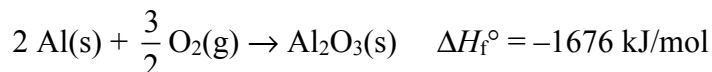
A thermochemical equation for the formation of aluminum oxide, $Al_2O_3(s)$ with the energy term embedded in the equation is:

$$2 \text{ Al(s)} + \frac{3}{2}O_2(g) \rightarrow Al_2O_3(s) + 1676 \text{ kJ}$$
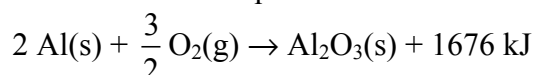
**29. (a)** Answers may vary. Sample answer: Hydrogen gas has a high enthalpy of combustion, releasing about 2.5 times the quantity of energy per gram than methane but the widespread use of hydrogen as a fuel has the limitations of high production cost and difficulty in storage and transport.

**(b)** Answers may vary. Sample answer: Possible discoveries that could make hydrogen a more viable alternative to fossil fuels include
• being able to generate hydrogen from another fuel when needed rather than relying on fueling station
• developing a very strong hydrogen fuel tank that can hold a great deal of compressed hydrogen safely

**30.** Fuel B is recommend for use in an airplane because for the same mass of fuel, fuel B releases more energy, 20 kJ/g, than fuel A during combustion.

**31. (a)** Answers may vary. Sample answer: Someone might choose to use electrical energy generated from a wind farm because the advances in wind turbine technology have provided significant improvements in efficiency. Wind energy is a more environmental friendly alternative since turbines provide clean renewable energy.

**(b)** Answers may vary. Sample answer: Someone might choose to not to use electrical energy generated from a wind farm because wind technology is only viable in regions with strong and consistent winds and that even though it is efficient, it is still not as efficient as the combustion of fossil fuels.

**Analysis and Application**

**32.** Answers may vary. Students' graphic organizers should include comparison between kinetic and potential energy, as well as concepts and real world examples of each type of energy.

**33.** Answers may vary. Students' diagrams should include what happens to the water molecules and any energy transfers that take place. Ideas could include: The process of liquid water turning into steam is an endothermic process. Thermal energy is the sum of the kinetic and potential energy. With more thermal energy, the kinetic energy of the water molecules increases and the temperature of water increases. At 100 °C, the potential energy of the system also increases during the phase change from liquid water to steam. The thermal energy lost by the system is transferred to the surroundings, causing an increase in temperature.

**34. (a)** In the balanced chemical equation, when 4 mol of water is formed, 1454 kJ of thermal energy is released. Calculate the quantity of energy released for 1 mol of water.

$$\frac{1452 \text{ kJ}}{4 \text{ mol}} = 363 \text{ kJ/mol}$$

So, the quantity of thermal energy released for each mole of water formed is 363 kJ.

**(b)** In the balanced chemical equation, when 2 mol of carbon dioxide is formed, 1454 kJ of thermal energy is released. Calculate the quantity of energy released for 1 mol of carbon dioxide.

$$\frac{1452 \text{ kJ}}{2 \text{ mol}} = 726 \text{ kJ/mol}$$

So, the quantity of thermal energy released for each mole of carbon dioxide formed is 726 kJ.

**(c)** In the balanced chemical equation, when 3 mol of oxygen is reacted, 1454 kJ of thermal energy is released. Calculate the quantity of energy released for 1 mol of oxygen.

$$\frac{1452 \text{ kJ}}{3 \text{ mol}} = 484 \text{ kJ/mol}$$

So, the quantity of thermal energy released for each mole of oxygen reacted is 484 kJ.

**(d)** In the balanced chemical equation, when 2 mol of methanol is used for combustion, 1454 kJ of thermal energy is released. Calculate the quantity of energy released for 1 mol of methanol.

$$\frac{1452 \text{ kJ}}{2 \text{ mol}} = 726 \text{ kJ/mol}$$

So, the molar enthalpy change, $\Delta H_{comb}$, of this reaction is $-726$ kJ/mol.

**35. (a) Given:** $m_{SiO_2(s)} = 1.00$ g; $\Delta T = 10.5$ °C; $c_{SiO_2(s)} = 0.835$ J/(g·°C)

**Required**: amount of energy absorbed, $q$

**Analysis:** $q = mc\Delta T$

**Solution:**

$q = mc\Delta T$

$$= (1.00 \text{ g})\left(\frac{0.835 \text{ J}}{\text{g} \cdot \text{°C}}\right)(10.5 \text{ °C})$$

$q = 8.77$ J

**Statement:** The amount of energy absorbed by 1.00 g of silicon dioxide, $SiO_2$(s), when its temperature is raised 10.5 °C is 8.77 kJ.

**(b) Given:** $m_{H_2O(l)} = 1.00$ g; $\Delta T = 10.5$ °C; $c_{H_2O(l)} = 4.18$ J/(g·°C)

**Required**: amount of energy absorbed, $q$

**Analysis:** $q = mc\Delta T$

**Solution:**

$q = mc\Delta T$

$$= (1.00 \text{ g})\left(\frac{4.18 \text{ J}}{\text{g} \cdot {}^\circ\text{C}}\right)(10.5 \text{ }{}^\circ\text{C})$$

$q = 43.9 \text{ J}$

**Statement:** The amount of energy absorbed by 1.00 g of water, $H_2O(l)$, when its temperature is raised 10.5 °C is 43.9 J.

**36. Given:** $V_{HCl(aq)} = 100.0 \text{ mL}$; $V_{NaOH(aq)} = 100.0 \text{ mL}$; $T_{initial} = 23.5 \text{ °C}$; $T_{final} = 30.6 \text{ °C}$

**Required**: quantity of energy released, $q$

**Analysis:** $q = mc\Delta T$

**Solution:**

**Step 1:** Determine the total volume of HCl(aq) and NaOH(aq), $V$.

$V = V_{HCl(aq)} + V_{NaOH(aq)}$

$= 100.0 \text{ mL} + 100.0 \text{ mL}$

$V = 200.0 \text{ mL}$

**Step 2:** Determine the total mass of the two solutions, $m$.

Since the solution contains dilute HCl(aq) and NaOH(aq), the density and heat capacity are assumed to be the same as those for water.

$m = Vd$

$$= 200.0 \text{ mL} \times \frac{1.00 \text{ g}}{1 \text{ mL}}$$

$m = 200.0 \text{ g}$

**Step 3:** Determine the change in temperature, $\Delta T$.

$\Delta T = T_{final} - T_{initial}$

$= 30.6 \text{ °C} - 23.5 \text{ °C}$

$\Delta T = 7.1 \text{ °C}$

**Step 4:** Calculate the quantity of energy released, $q$.

$q = mc\Delta T$

$$= (200.0 \text{ g})\left(\frac{4.18 \text{ J}}{\text{g} \cdot {}^\circ\text{C}}\right)(7.1 \text{ }{}^\circ\text{C})$$

$q = 5936 \text{ J}$ (2 extra digits carried)

**Statement:** The quantity of energy released in the reaction is 5900 J, or 5.9 kJ.

**(b) Given:** $V_{HCl(aq)} = 100.0 \text{ mL}$; $q = 5936 \text{ J}$; [HCl(aq)] = 1.0 mol/L

**Required**: energy released per mole of hydrochloric acid

**Analysis:** $\dfrac{q}{n_{HCl(aq)}}$

**Solution:**

**Step 1:** Determine the amount of hydrochloric acid in 100.0 mL of 1.0 mol/L solution.

$$n_{HCl(aq)} = 100.0 \text{ mL} \times \frac{1.0 \text{ mol}}{1 \text{ L}} \times \frac{1 \text{ L}}{1000 \text{ mL}}$$

$n_{HCl(aq)} = 0.010 \text{ mol}$

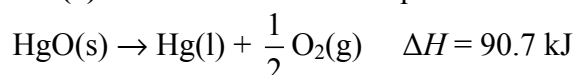**Step 2:** Calculate energy per mole of hydrochloric acid released in the reaction.

$$\frac{q}{n_{HCl(aq)}} = \frac{5936 \text{ J}}{0.010 \text{ mol}}$$

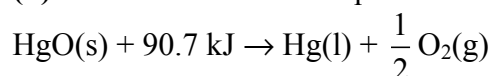$$= \frac{5936 \text{ J}}{0.010 \text{ mol}} \times \frac{1 \text{ kJ}}{1000 \text{ J}}$$

$$\frac{q}{n_{HCl(aq)}} = 59 \text{ kJ/mol}$$

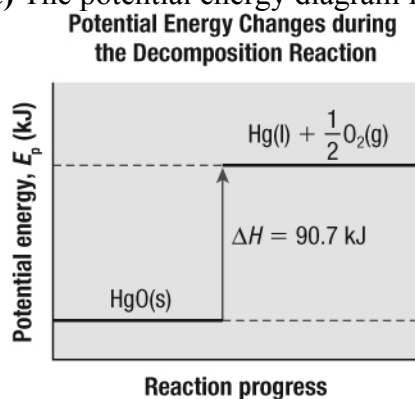**Statement:** The quantity of energy per mole of hydrochloric acid released in the reaction is 59 kJ/mol.

**37. (a)** The thermochemical equation for the reaction with the energy term separate is:

$$HgO(s) \rightarrow Hg(l) + \frac{1}{2} O_2(g) \quad \Delta H = 90.7 \text{ kJ}$$

**(b)** The thermochemical equation for the reaction with the energy term embedded is:

$$HgO(s) + 90.7 \text{ kJ} \rightarrow Hg(l) + \frac{1}{2} O_2(g)$$

**(c)** The potential energy diagram for the reaction is:



**38. Given:** $m_{NH_4NO_3(s)} = 1.60 \text{ g}$; $m_{H_2O(l)} = 75.0 \text{ g}$; $T_{initial} = 25.00 \text{ °C}$; $T_{final} = 23.34 \text{ °C}$

**Required:** molar enthalpy of dissolution, $\Delta H_{sol}$

**Analysis:** $q = mc\Delta T$; $\Delta H = n\Delta H_{sol}$

**Solution:**

**Step 1:** Determine the change in temperature, $\Delta T$.

$\Delta T = T_{final} - T_{initial}$

$\quad = 23.34 \text{ °C} - 25.00 \text{ °C}$

$\Delta T = 1.66 \text{ °C}$

**Step 2:** Determine the quantity of thermal energy absorbed by the water, $q$.

$q = mc\Delta T$

$$= (75.0 \text{ g}) \left( \frac{4.18 \text{ J}}{\text{g} \cdot \text{°C}} \right) (1.66 \text{ °C})$$

$q = 520.41 \text{ J}$ (2 extra digits carried)

This is the quantity of energy released, $\Delta H$, when 1.60 g of $NH_4NO_3(s)$ is dissolved.

**Step 3:** Convert mass of $NH_4NO_3(s)$ into amount, $n_{NH_4NO_3(s)}$.

$M_{NH_4NO_3(s)} = 80.06$ g/mol

$n_{NH_4NO_3(s)} = \dfrac{m_{NH_4NO_3(s)}}{M_{NH_4NO_3(s)}}$

$= \dfrac{1.60 \cancel{g}}{80.06 \cancel{g}/mol}$

$n_{NH_4NO_3(s)} = 0.019\ 985$ mol (2 extra digits carried)

**Step 4:** Rearrange $\Delta H = n\Delta H_{sol}$ to solve for the molar enthalpy of dissolution, $\Delta H_{sol}$.
Since the drop in temperature indicates that energy is absorbed, $\Delta H$ for the reaction is positive;
$\Delta H = 520.41$ J.

$\Delta H = n_{NH_4NO_3(s)}\Delta H_{sol}$

$\Delta H_{sol} = \dfrac{\Delta H}{n_{NH_4NO_3(s)}}$

$= \dfrac{520.41\ J}{0.019\ 985\ mol}$

$\Delta H_{sol} = 26\ 000$ J/mol

**Statement:** The enthalpy change for the dissolution of ammonium nitrate is $26.0 \times 10^3$ J/mol, or 26.0 kJ/mol.

**39. Given:** $m_{K(s)} = 60.0$ g; enthalpy change of given reaction $= -160$ kJ

**Required**: enthalpy change for 60.0 g of K(s), $\Delta H$

**Analysis:** $\Delta H = n\Delta H_r$

**Solution:**

**Step 1:** Calculate the thermal energy released by 1 mol of K(s).
From the equation, 160 kJ of energy is released by the reaction of 2 mol of K(s).

$\Delta H_r = \dfrac{-160\ kJ}{2\ mol}$

$\Delta H_r = -80$ kJ/mol

**Step 2:** Calculate the amount of potassium in 60.0 g, $n_{K(s)}$.

$M_{K(s)} = 39.10$ g/mol

$n_{K(s)} = \dfrac{m_{K(s)}}{M_{K(s)}}$

$= \dfrac{60.0 \cancel{g}}{39.10 \cancel{g}/mol}$

$n_{K(s)} = 1.5345$ mol (2 extra digits carried)

**Step 3:** Solve for the change in enthalpy, $\Delta H$.

$$\Delta H = n_{K(s)}\Delta H_r$$
$$= (1.5345 \ \cancel{\text{mol}})\left(\frac{-80 \text{ kJ}}{1 \ \cancel{\text{mol}}}\right)$$
$$\Delta H = -120 \text{ kJ}$$

**Statement:** The thermal energy released when 60.0 g of potassium metal reacts is 120 kJ.

**40. (a)** As ethanol changes state from a solid to a liquid, the molecules absorb energy from the surroundings. The molecules enter a higher energy state and begin to vibrate faster. This causes the molecules to be able to flow and slide over each other. As ethanol changes state from a liquid to a vapour, the molecules absorb energy from the surroundings to enter an even higher energy state. The molecules move much faster and are much farther apart.

**(b)** For ethanol, $\Delta H_{vap} = +38.6 \text{ kJ/mol}$.

To calculate the energy required to vaporize 2.0 mol of ethanol, use $\Delta H = n\Delta H_{vap}$.

$$\Delta H = n\Delta H_{vap}$$
$$= 2.0 \ \cancel{\text{mol}} \times \frac{38.6 \text{ kJ}}{1 \ \cancel{\text{mol}}}$$
$$\Delta H = 77.2 \text{ kJ}$$

So, the energy required to vaporize 2.0 mol of ethanol is 77.2 kJ.

**(c)** Fusion is the reverse of melting.

$$\Delta H_{melt} = -\Delta H_{fus}$$
$$\Delta H_{melt} = 26.1 \text{ kJ/mol}$$

To calculate the energy required to melt 0.5 mol of solid ethanol, use $\Delta H = n\Delta H_{melt}$.

$$\Delta H = n\Delta H_{vap}$$
$$= 0.5 \ \cancel{\text{mol}} \times \frac{26.1 \text{ kJ}}{1 \ \cancel{\text{mol}}}$$
$$\Delta H = 13.1 \text{ kJ}$$

So, the energy required to melt 0.5 mol of solid ethanol is 13.1 kJ.

**41. (a)** Experiment 1:

**Given:** $m_{CaCO_3(s)} = 4.2 \text{ g}$; $m_{HCl(aq)} = 170.2 \text{ g}$; $\Delta T_{initial} = 24.0 \text{ °C}$; $\Delta T_{final} = 26.0 \text{ °C}$; $c = 4.18 \text{ J/(g·°C)}$

**Required:** $\Delta H_1$

**Analysis:** $q = mc\Delta T$; $\Delta H = -q_{surroundings}$

**Solution:**

**Step 1:** Determine the change in temperature, $\Delta T$.

$$\Delta T = T_{final} - T_{initial}$$
$$= 26.0 \text{ °C} - 24.0 \text{ °C}$$
$$\Delta T = 2.0 \text{ °C}$$

**Step 2:** Determine the quantity of thermal energy absorbed by the solution, $q$.

$q = mc\Delta T$

$$= (170.2 \text{ g}) \left( \frac{4.18 \text{ J}}{\text{g} \cdot {}^{\circ}\text{C}} \right) (2.0 \text{ }{}^{\circ}\text{C})$$

$q = 1423$ J (2 extra digits carried)

**Step 3:** Calculate $\Delta H_1$.

$\Delta H_1 = -q_{\text{surroundings}}$

$\Delta H_1 = -1423$ J (2 extra digits carried)

**Step 4:** Change the amount of $CaCO_3(s)$ in 4.2 g of the substance.

$M_{CaCO_3(s)} = 100.09$ g/mol

$$n_{CaCO_3(s)} = \frac{m_{CaCO_3(s)}}{M_{CaCO_3(s)}}$$

$$= \frac{4.2 \text{ g}}{100.09 \text{ g/mol}}$$

$n_{CaCO_3(s)} = 4.196 \times 10^{-2}$ mol (2 extra digits carried)

**Step 5:** Calculate the enthalpy change, $\Delta H$, of the reaction for 1 mol of $CaCO_3(s)$.

Use $\Delta H_1 = n\Delta H$

$$\Delta H = \frac{\Delta H_1}{n_{CaCO_3(s)}}$$

$$= \frac{-1423 \text{ J}}{4.196 \times 10^{-2} \text{ mol}}$$

$\Delta H = -34$ kJ/mol (2 extra digits carried)

**Statement:** The enthalpy change of reaction 1 is –34 kJ/mol.

Experiment 2:

**Given:** $m_{CaCO_3(s)} = 4.7$ g; $m_{HCl(aq)} = 155.5$ g; $\Delta T_{\text{initial}} = 23.8 \text{ °C}$; $\Delta T_{\text{final}} = 30.8 \text{ °C}$;

$c = 4.18$ J/(g·°C)

**Required:** $\Delta H_2$

**Analysis:** $q = mc\Delta T$; $\Delta H = -q_{\text{surroundings}}$

**Solution:**

**Step 1:** Determine the change in temperature, $\Delta T$.

$\Delta T = T_{\text{final}} - T_{\text{initial}}$

$= 30.8 \text{ °C} - 23.8 \text{ °C}$

$\Delta T = 7.0 \text{ °C}$

**Step 2:** Determine the quantity of thermal energy absorbed by the solution, $q$.

$q = mc\Delta T$

$$= (155.5 \text{ g}) \left( \frac{4.18 \text{ J}}{\text{g} \cdot {}^{\circ}\text{C}} \right) (7.0 \text{ }{}^{\circ}\text{C})$$

$q = 4550$ J (2 extra digits carried)

**Step 3:** Calculate $\Delta H_2$.

$\Delta H_2 = -q_{surroundings}$

$\Delta H_2 = -4550$ J (2 extra digits carried)

**Step 4:** Change the amount of CaCO(s) in 4.7 g of the substance.

$M_{CaO(s)} = 56.08$ g/mol

$n_{CaO(s)} = \dfrac{m_{CaO(s)}}{M_{CaO(s)}}$

$= \dfrac{4.7 \text{ g}}{56.08 \text{ g/mol}}$

$n_{CaO(s)} = 8.381 \times 10^{-2}$ mol (2 extra digits carried)

**Step 5:** Calculate the enthalpy change, $\Delta H$, of the reaction for 1 mol of CaCO(s).

Use $\Delta H_2 = n\Delta H$

$\Delta H = \dfrac{\Delta H_2}{n_{CaO(s)}}$
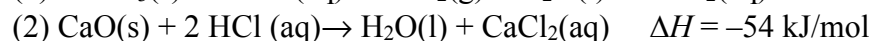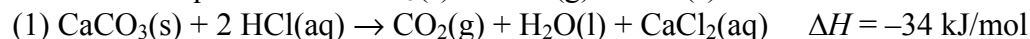
$= \dfrac{-4550 \text{ J}}{8.381 \times 10^{-2} \text{ mol}}$
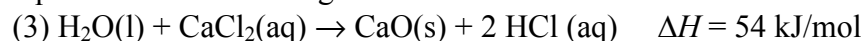
$\Delta H = -54$ kJ/mol

**Statement:** The enthalpy change of reaction 2 is –54 kJ/mol.

**(b) Solution:**

**Step 1:** Label the equations.

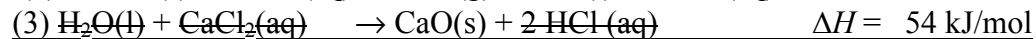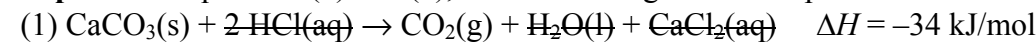The desired equation is: $CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

(1) $CaCO_3(s) + 2 HCl(aq) \rightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$    $\Delta H = -34$ kJ/mol

(2) $CaO(s) + 2 HCl (aq) \rightarrow H_2O(l) + CaCl_2(aq)$    $\Delta H = -54$ kJ/mol

**Step 2:** Reverse equation (2) so that reactants and products are on the same side as in the desired equation. Reverse the sign of $\Delta H$.

(3) $H_2O(l) + CaCl_2(aq) \rightarrow CaO(s) + 2 HCl (aq)$    $\Delta H = 54$ kJ/mol

**Step 3:** Add equations (1) and (3), and their changes in enthalpies.

(1) $CaCO_3(s) + \cancel{2 HCl(aq)} \rightarrow CO_2(g) + \cancel{H_2O(l)} + \cancel{CaCl_2(aq)}$    $\Delta H = -34$ kJ/mol

(3) $\cancel{H_2O(l)} + \cancel{CaCl_2(aq)} \rightarrow CaO(s) + \cancel{2 HCl (aq)}$    $\Delta H = 54$ kJ/mol
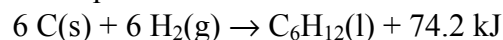
$CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$    $\Delta H = 20$ kJ/mol

**Statement:** The enthalpy change of the desired reaction is 20 kJ/mol.

**42. (a)** The equation for the formation reaction for liquid cyclohexane, $C_6H_{12}(l)$, is

$6 C(s) + 6 H_2(g) \rightarrow C_6H_{12}(l) + 156$ kJ
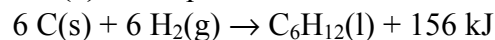
The equation for the formation reaction for liquid hex-1-ene, $C_6H_{12}(l)$, is

$6 C(s) + 6 H_2(g) \rightarrow C_6H_{12}(l) + 74.2$ kJ

**(b) Given:** $\Delta H_f{}^{\circ}{}_{\text{cyclohexane}} = -156 \text{ kJ/mol}$; $\Delta H_f{}^{\circ}{}_{\text{hex-1-ene}} = -74.2 \text{ kJ/mol}$

$\Delta H_f{}^{\circ}{}_{CO_2(g)} = -393.5 \text{ kJ/mol}$; $\Delta H_f{}^{\circ}{}_{H_2O(l)} = -285.8 \text{ kJ/mol}$; $\Delta H_f{}^{\circ}{}_{O_2(g)} = 0 \text{ kJ/mol}$

**Required:** $\Delta H_c{}^{\circ}{}_{\text{cyclohexane}}$; $\Delta H_c{}^{\circ}{}_{\text{hex-1-ene}}$

**Analysis:** $\Delta H_r{}^{\circ} = \sum n_{\text{products}} \times \Delta H_{\text{products}} - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}$
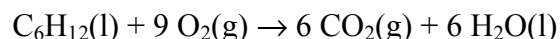
Since $O_2(g)$ is in its standard state, the equation can be written as

$\Delta H_c{}^{\circ}{}_{\text{cyclohexane}} = [n_{CO_2(g)} \times \Delta H_f{}^{\circ}{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^{\circ}{}_{H_2O(l)}] - n_{\text{cyclohexane}} \times \Delta H_f{}^{\circ}{}_{\text{cyclohexane}}$

$\Delta H_c{}^{\circ}{}_{\text{hex-1-ene}} = [n_{CO_2(g)} \times \Delta H_f{}^{\circ}{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^{\circ}{}_{H_2O(l)}] - n_{\text{hex-1-ene}} \times \Delta H_f{}^{\circ}{}_{\text{hex-1-ene}}$

**Solution:**

**Step 1:** Write a balanced chemical equation for the combustion of liquid cyclohexane or hex-1-ene so that the substance has a coefficient of 1.

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

**Step 2:** Substitute the appropriate values for standard enthalpy into the equation for each substance and solve.

$\Delta H_c{}^{\circ}{}_{\text{cyclohexane}} = [n_{CO_2(g)} \times \Delta H_f{}^{\circ}{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^{\circ}{}_{H_2O(l)}] - n_{\text{cyclohexane}} \times \Delta H_f{}^{\circ}{}_{\text{cyclohexane}}$

$\phantom{\Delta H_c{}^{\circ}{}_{\text{cyclohexane}}} = [6(-393.5 \text{ kJ}) + 6(-285.8 \text{ kJ})] - (-156 \text{ kJ})$

$\phantom{\Delta H_c{}^{\circ}{}_{\text{cyclohexane}}} = -2361 \text{ kJ} + 1714.8 \text{ kJ} + 156 \text{ kJ}$

$\Delta H_c{}^{\circ}{}_{\text{cyclohexane}} = -3920 \text{ kJ}$

$\Delta H_c{}^{\circ}{}_{\text{hex-1-ene}} = [n_{CO_2(g)} \times \Delta H_f{}^{\circ}{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^{\circ}{}_{H_2O(l)}] - n_{\text{hex-1-ene}} \times \Delta H_f{}^{\circ}{}_{\text{hex-1-ene}}$

$\phantom{\Delta H_c{}^{\circ}{}_{\text{hex-1-ene}}} = [6(-393.5 \text{ kJ}) + 6(-285.8 \text{ kJ})] - (-74.2 \text{ kJ})$

$\phantom{\Delta H_c{}^{\circ}{}_{\text{hex-1-ene}}} = -2361.0 \text{ kJ} - 1714.8 \text{ kJ} + 74.2 \text{ kJ}$

$\Delta H_c{}^{\circ}{}_{\text{hex-1-ene}} = -4001.6 \text{ kJ}$

**Statement:** The enthalpy change for the combustion of liquid cyclohexane is –3920 kJ and that for liquid 1-hexene is –4001.6 kJ.

**(c)** Cyclohexane releases a similar amount of energy as hex-1-ene upon combustion. This indicates that the ring structure is similar to the structure with a double than to the structure of the corresponding alkane.

**43. (a)** There are 2 C–H bonds and 2 C–Cl bonds in $CH_2Cl_2$.

**Given:** $n_{\text{C-H}} = 2 \text{ mol}$; $D_{\text{C-H}} = 413 \text{ kJ/mol}$; $n_{\text{C-Cl}} = 2 \text{ mol}$; $D_{\text{C-Cl}} = 339 \text{ kJ/mol}$

**Required:** $\Delta H$

**Analysis:** $\Delta H = \sum n \times D_{\text{bonds broken}}$

**Solution:**

$\Delta H = \sum n \times D_{\text{bonds broken}}$

$\phantom{\Delta H} = 2 \text{ mol} \times D_{\text{C-H}} + 2 \text{ mol} \times D_{\text{C-Cl}}$

$\phantom{\Delta H} = 2 \, \cancel{\text{mol}} \times \dfrac{413 \text{ kJ}}{\cancel{\text{mol}}} + 2 \, \cancel{\text{mol}} \times \dfrac{339 \text{ kJ}}{\cancel{\text{mol}}}$

$\Delta H = 1504 \text{ kJ}$

**Statement:** The atomization energy for $CH_2Cl_2$ is 1504 kJ/mol.

**(b)** There are 3 C–H bonds and 1 C–Br bond in $CH_3Br$.

**Given:** $n_{C-H} = 3 \text{ mol}$; $D_{C-H} = 413 \text{ kJ/mol}$; $n_{C-Br} = 1 \text{ mol}$; $D_{C-Br} = 276 \text{ kJ/mol}$

**Required:** $\Delta H$

**Analysis:** $\Delta H = \sum n \times D_{\text{bonds broken}}$

**Solution:**

$$\Delta H = \sum n \times D_{\text{bonds broken}}$$
$$= 3 \text{ mol} \times D_{C-H} + 1 \text{ mol} \times D_{C-Br}$$
$$= 3 \text{ mol} \times \frac{413 \text{ kJ}}{\text{mol}} + 1 \text{ mol} \times \frac{276 \text{ kJ}}{\text{mol}}$$

$\Delta H = 1515 \text{ kJ}$

**Statement:** The atomization energy for $CH_3Br$ is 1515 kJ/mol.

**(c)** There are 2 C–H bonds and 1 C≡C bond in $C_2H_2$.

**Given:** $n_{C-H} = 2 \text{ mol}$; $D_{C-H} = 413 \text{ kJ/mol}$; $n_{C\equiv C} = 1 \text{ mol}$; $D_{C\equiv C} = 839 \text{ kJ/mol}$

**Required:** $\Delta H$

**Analysis:** $\Delta H = \sum n \times D_{\text{bonds broken}}$

**Solution:**

$$\Delta H = \sum n \times D_{\text{bonds broken}}$$
$$= 2 \text{ mol} \times D_{C-H} + 1 \text{ mol} \times D_{C\equiv C}$$
$$= 2 \text{ mol} \times \frac{413 \text{ kJ}}{\text{mol}} + 1 \text{ mol} \times \frac{839 \text{ kJ}}{\text{mol}}$$

$\Delta H = 1665 \text{ kJ}$

**Statement:** The atomization energy for $C_2H_2$ is 1665 kJ/mol.

**(d)** There are 4 C–H bonds and 1 C=C bond in $CH_2{=}CH_2$.

**Given:** $n_{C-H} = 4 \text{ mol}$; $D_{C-H} = 413 \text{ kJ/mol}$; $n_{C=C} = 1 \text{ mol}$; $D_{C=C} = 614 \text{ kJ/mol}$
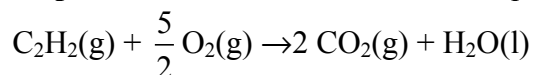
**Required:** $\Delta H$

**Analysis:** $\Delta H = \sum n \times D_{\text{bonds broken}}$

**Solution:**

$$\Delta H = \sum n \times D_{\text{bonds broken}}$$
$$= 4 \text{ mol} \times D_{C-H} + 1 \text{ mol} \times D_{C=C}$$
$$= 4 \text{ mol} \times \frac{413 \text{ kJ}}{\text{mol}} + 1 \text{ mol} \times \frac{614 \text{ kJ}}{\text{mol}}$$

$\Delta H = 2266 \text{ kJ}$

**Statement:** The atomization energy for $CH_2{=}CH_2$ is 2266 kJ/mol.

**44. (a) Solution:**
**Step 1:** Write the balanced chemical equation for the combustion 1 mol of acetylene gas.

$$C_2H_2(g) + \frac{5}{2} O_2(g) \rightarrow 2\ CO_2(g) + H_2O(l)$$

Determine the bonding of each substance.
**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.
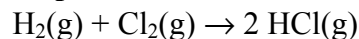
|  | Substance | Number of bonds per mole ($n_{\text{substance}}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $C_2H_2$ | 2 mol C–H bonds | 2 mol | 413 kJ/mol |
|  |  | 1 mol C≡C bonds | 1 mol | 839 kJ/mol |
|  | $O_2$ | 1 mol O=O bonds | $\frac{5}{2}$ mol | 495 kJ/mol |
| products | $CO_2$ | 2 mol C=O bonds | 4 mol | 799 kJ/mol |
|  | $H_2O$ | 2 mol H–O bonds | 2 mol | 467 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 2\ \text{mol} \times D_{\text{C–H}} + 1\ \text{mol} \times D_{\text{C≡C}} + \frac{5}{2}\ \text{mol} \times D_{\text{O=O}} -$$

$$\left( 4\ \text{mol} \times D_{\text{C=O}} + 2\ \text{mol} \times D_{\text{H–O}} \right)$$

$$= \left( 2\ \cancel{\text{mol}} \times \frac{413\ \text{kJ}}{\cancel{\text{mol}}} \right) + \left( 1\ \cancel{\text{mol}} \times \frac{839\ \text{kJ}}{\cancel{\text{mol}}} \right) + \left( \frac{5}{2}\ \cancel{\text{mol}} \times \frac{495\ \text{kJ}}{\cancel{\text{mol}}} \right) -$$

$$\left[ \left( 4\ \cancel{\text{mol}} \times \frac{799\ \text{kJ}}{\cancel{\text{mol}}} \right) + \left( 2\ \cancel{\text{mol}} \times \frac{467\ \text{kJ}}{\cancel{\text{mol}}} \right) \right]$$

$$= 826\ \text{kJ} + 839\ \text{kJ} + 1237.5\ \text{kJ} - (3196\ \text{kJ} + 934\ \text{kJ})$$

$\Delta H = -1228\ \text{kJ}$
**Statement:** Since enthalpy change, $\Delta H$, of the reaction is negative, the complete combustion of acetylene gas is an exothermic reaction.
**(b)** Answer may vary. Sample answer: Bond energy values can be used to calculate approximate enthalpy changes, $\Delta H$, for reactions. For bonds to be broken, energy must be added, and when bonds are formed, energy is released. First, determine how many of each type of bond must be broken in the reactants. Next, determine how many of each type of bonds form in the products. Finally, use bond dissociation energies to calculate the total energy required to break the reactant bonds, followed by the total energy released by the formation of product bonds. The energy change, $\Delta H$, of the reaction is the difference between these two sums.

**45. (a) Solution:**
**Step 1:** Use the balanced chemical equation:
$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$
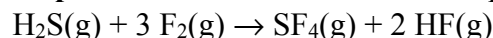Determine the bonding of each substance.
**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

|  | Substance | Number of bonds per mole ($n_{substance}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $H_2$ | 1 mol H–H bonds | 1 mol | 432 kJ/mol |
|  | $Cl_2$ | 1 mol Cl–Cl bonds | 1 mol | 239 kJ/mol |
| products | HCl | 1 mol H–Cl bonds | 2 mol | 427 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.

$$\Delta H = \sum n \times D_{bonds\ broken} - \sum n \times D_{bonds\ formed}$$

$$= 1\ mol \times D_{H-H} + 1\ mol \times D_{Cl-Cl} - \left(2\ mol \times D_{H-Cl}\right)$$

$$= \left(1\ mol \times \frac{432\ kJ}{mol}\right) + \left(1\ mol \times \frac{239\ kJ}{mol}\right) - \left(2\ mol \times \frac{427\ kJ}{mol}\right)$$

$$= 432\ kJ + 239\ kJ - (854\ kJ)$$

$\Delta H = -183\ kJ$

**Statement:** The enthalpy change, $\Delta H$, for the reaction is –183 kJ. Since $\Delta H$ is negative, the reaction is exothermic.

**(b) Solution:**
**Step 1:** Use the balanced chemical equation:
$H_2S(g) + 3\ F_2(g) \rightarrow SF_4(g) + 2\ HF(g)$
Determine the bonding of each substance.
**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

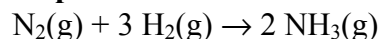|  | Substance | Number of bonds per mole ($n_{substance}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $H_2S$ | 2 mol S–H bonds | 2 mol | 347 kJ/mol |
|  | $F_2$ | 1 mol F–F bonds | 3 mol | 154 kJ/mol |
| products | $SF_4$ | 4 mol S–F bonds | 4 mol | 327 kJ/mol |
|  | HF | 1 mol H–F bonds | 2 mol | 565 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 2 \text{ mol} \times D_{\text{S–H}} + 3 \text{ mol} \times D_{\text{F–F}} - (4 \text{ mol} \times D_{\text{S–F}} + 2 \text{ mol} \times D_{\text{H–F}})$$

$$= \left( 2 \text{ mol} \times \frac{347 \text{ kJ}}{\text{mol}} \right) + \left( 3 \text{ mol} \times \frac{154 \text{ kJ}}{\text{mol}} \right) -$$

$$\left[ \left( 4 \text{ mol} \times \frac{327 \text{ kJ}}{\text{mol}} \right) + \left( 2 \text{ mol} \times \frac{565 \text{ kJ}}{\text{mol}} \right) \right]$$

$$= 694 \text{ kJ} + 462 \text{ kJ} - (1308 \text{ kJ} + 1130 \text{ kJ})$$

$\Delta H = -1282 \text{ kJ}$

**Statement:** The enthalpy change, $\Delta H$, for the reaction is $-1282$ kJ. Since $\Delta H$ is negative, the reaction is exothermic.

**(c) Solution:**

**Step 1:** Use the balanced chemical equation:

$N_2(g) + 3 H_2(g) \rightarrow 2 NH_3(g)$

Determine the bonding of each substance.

**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

| | Substance | Number of bonds per mole ($n_{\text{substance}}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $N_2$ | 1 mol N≡N bonds | 1 mol | 941 kJ/mol |
| | $H_2$ | 1 mol H–H bonds | 3 mol | 432 kJ/mol |
| products | $NH_3$ | 3 mol N–H bonds | 6 mol | 391 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.
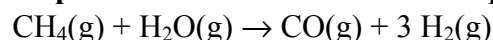
$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 1 \text{ mol} \times D_{\text{N≡N}} + 3 \text{ mol} \times D_{\text{H–H}} - (6 \text{ mol} \times D_{\text{N–H}})$$

$$= \left( 1 \text{ mol} \times \frac{941 \text{ kJ}}{\text{mol}} \right) + \left( 3 \text{ mol} \times \frac{432 \text{ kJ}}{\text{mol}} \right) - \left( 6 \text{ mol} \times \frac{391 \text{ kJ}}{\text{mol}} \right)$$

$$= 941 \text{ kJ} + 1296 \text{ kJ} - (2346 \text{ kJ})$$

$\Delta H = -109 \text{ kJ}$

**Statement:** The enthalpy change, $\Delta H$, for the reaction is $-109$ kJ. Since $\Delta H$ is negative, the reaction is exothermic.

**(d) Solution:**

**Step 1:** Write the equation that represents the complete combustion of propene.
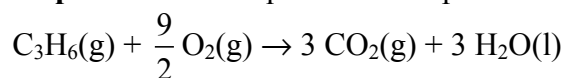
$C_3H_6(g) + \dfrac{9}{2} O_2(g) \rightarrow 3 CO_2(g) + 3 H_2O(l)$

Determine the bonding of each substance.

**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

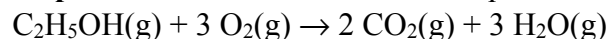| | Substance | Number of bonds per mole ($n_{substance}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $C_3H_6$ | 6 mol C–H bonds | 6 mol | 413 kJ/mol |
| | | 1 mol C–C bonds | 1 mol | 347 kJ/mol |
| | | 1 mol C=C bonds | 1 mol | 614 kJ/mol |
| | $O_2$ | 1 mol O=O bonds | $\frac{9}{2}$ mol | 495 kJ/mol |
| products | $CO_2$ | 2 mol C=O bonds | 6 mol | 799 kJ/mol |
| | $H_2O$ | 2 mol H–O bonds | 6 mol | 467 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 6 \text{ mol} \times D_{\text{C–H}} + 1 \text{ mol} \times D_{\text{C–C}} + 1 \text{ mol} \times D_{\text{C=C}} + \frac{9}{2} \text{ mol} \times D_{\text{O=O}} -$$

$$\left( 6 \text{ mol} \times D_{\text{C=O}} + 6 \text{ mol} \times D_{\text{H–O}} \right)$$

$$= \left( 6 \text{ mol} \times \frac{413 \text{ kJ}}{\text{mol}} \right) + \left( 1 \text{ mol} \times \frac{347 \text{ kJ}}{\text{mol}} \right) + \left( 1 \text{ mol} \times \frac{614 \text{ kJ}}{\text{mol}} \right) + \left( \frac{9}{2} \text{ mol} \times \frac{495 \text{ kJ}}{\text{mol}} \right) -$$

$$\left[ \left( 6 \text{ mol} \times \frac{799 \text{ kJ}}{\text{mol}} \right) + \left( 6 \text{ mol} \times \frac{467 \text{ kJ}}{\text{mol}} \right) \right]$$

$$= 2478 \text{ kJ} + 347 \text{ kJ} + 614 \text{ kJ} + 2227.5 \text{ kJ} - (4794 \text{ kJ} + 2802 \text{ kJ})$$

$$\Delta H = -1930 \text{ kJ}$$

**Statement:** The enthalpy change, $\Delta H$, for the reaction is $-1930$ kJ. Since $\Delta H$ is negative, the reaction is exothermic.

**46. Solution:**
**Step 1:** Use the balanced chemical equation:
$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3 H_2(g)$
Determine the bonding of each substance.
**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

| | Substance | Number of bonds per mole ($n_{substance}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $CH_4$ | 4 mol C–H bonds | 4 mol | 413 kJ/mol |
| | $H_2O$ | 2 mol H–O bonds | 2 mol | 467 kJ/mol |
| products | CO | 1 mol C≡O bonds | 1 mol | 1072 kJ/mol |
| | $H_2$ | 1 mol H–H bonds | 3 mol | 432 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 4 \text{ mol} \times D_{\text{C–H}} + 2 \text{ mol} \times D_{\text{H–O}} - (1 \text{ mol} \times D_{\text{C}\equiv\text{O}} + 3 \text{ mol} \times D_{\text{H–H}})$$

$$= \left(4 \text{ mol} \times \frac{413 \text{ kJ}}{\text{mol}}\right) + \left(2 \text{ mol} \times \frac{467 \text{ kJ}}{\text{mol}}\right) - \left[\left(1 \text{ mol} \times \frac{1072 \text{ kJ}}{\text{mol}}\right) + \left(3 \text{ mol} \times \frac{432 \text{ kJ}}{\text{mol}}\right)\right]$$

$$= 1652 \text{ kJ} + 934 \text{ kJ} - (1072 \text{ kJ} + 1296 \text{ kJ})$$

$\Delta H = 218 \text{ kJ}$

**Statement:** The enthalpy change, $\Delta H$, for the reaction is 218 kJ.

**47. (a) Solution:**

**Step 1:** Use the balanced chemical equation:

$C_2H_5OH(g) + 3 \ O_2(g) \rightarrow 2 \ CO_2(g) + 3 \ H_2O(g)$

Determine the bonding of each substance.

**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

|  | Substance | Number of bonds per mole ($n_{\text{substance}}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $C_2H_5OH$ | 5 mol C–H bonds | 5 mol | 413 kJ/mol |
|  |  | 1 mol C–C bonds | 1 mol | 347 kJ/mol |
|  |  | 1 mol C–O bonds | 1 mol | 358 kJ/mol |
|  |  | 1 mol H–O bonds | 1 mol | 467 kJ/mol |
|  | $O_2$ | 1 mol O=O bonds | 3 mol | 495 kJ/mol |
| products | $CO_2$ | 2 mol C=O bonds | 4 mol | 799 kJ/mol |
|  | $H_2O$ | 2 mol H–O bonds | 6 mol | 467 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 5 \text{ mol} \times D_{\text{C–H}} + 1 \text{ mol} \times D_{\text{C–C}} + 1 \text{ mol} \times D_{\text{C–O}} + 1 \text{ mol} \times D_{\text{H–O}} + 3 \text{ mol} \times D_{\text{O=O}} -$$

$$(4 \text{ mol} \times D_{\text{C=O}} + 6 \text{ mol} \times D_{\text{H–O}})$$

$$= \left(5 \text{ mol} \times \frac{413 \text{ kJ}}{\text{mol}}\right) + \left(1 \text{ mol} \times \frac{347 \text{ kJ}}{\text{mol}}\right) + \left(1 \text{ mol} \times \frac{358 \text{ kJ}}{\text{mol}}\right) + \left(1 \text{ mol} \times \frac{467 \text{ kJ}}{\text{mol}}\right) +$$

$$\left(3 \text{ mol} \times \frac{495 \text{ kJ}}{\text{mol}}\right) - \left[\left(4 \text{ mol} \times \frac{799 \text{ kJ}}{\text{mol}}\right) + \left(6 \text{ mol} \times \frac{467 \text{ kJ}}{\text{mol}}\right)\right]$$

$$= 2065 \text{ kJ} + 347 \text{ kJ} + 358 \text{ kJ} + 467 \text{ kJ} + 1485 \text{ kJ} - (3196 \text{ kJ} + 2802 \text{ kJ})$$

$\Delta H = -1276 \text{ kJ}$

**Step 4:** Calculate the enthalpy change, $\Delta H_{\text{per gram}}$, of the reaction for 1.00 g of ethanol.
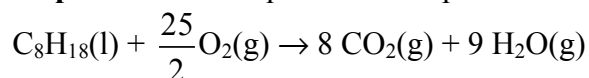
$M_{\text{C}_2\text{H}_5\text{OH(g)}} = 46.08$ g/mol

$$\Delta H_{\text{per gram}} = \frac{\Delta H}{M_{\text{C}_2\text{H}_5\text{OH(g)}}}$$

$$= \frac{-1276 \text{ kJ/mol}}{46.08 \text{ g/mol}}$$

$$\Delta H = -27.7 \text{ kJ/g}$$

**Statement:** The enthalpy of combustion per gram of ethanol is –27.69 kJ.

**(b) Solution:**

**Step 1:** Write the equation that represents the combustion reaction of octane.

$$\text{C}_8\text{H}_{18}(l) + \frac{25}{2}\text{O}_2(g) \rightarrow 8\ \text{CO}_2(g) + 9\ \text{H}_2\text{O}(g)$$

Determine the bonding of each substance.

**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.

|  | Substance | Number of bonds per mole ($n_{\text{substance}}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $\text{C}_8\text{H}_{18}$ | 18 mol C–H bonds | 18 mol | 413 kJ/mol |
|  |  | 7 mol C–C bonds | 7 mol | 347 kJ/mol |
|  | $\text{O}_2$ | 1 mol O=O bonds | $\frac{25}{2}$ mol | 495 kJ/mol |
| products | $\text{CO}_2$ | 2 mol C=O bonds | 16 mol | 799 kJ/mol |
|  | $\text{H}_2\text{O}$ | 2 mol H–O bonds | 18 mol | 467 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.

$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$

$$= 18 \text{ mol} \times D_{\text{C–H}} + 7 \text{ mol} \times D_{\text{C–C}} + \frac{25}{2} \text{ mol} \times D_{\text{O=O}} -$$

$$\left(16 \text{ mol} \times D_{\text{C=O}} + 18 \text{ mol} \times D_{\text{H–O}}\right)$$

$$= \left(18 \text{ mol} \times \frac{413 \text{ kJ}}{\text{mol}}\right) + \left(7 \text{ mol} \times \frac{347 \text{ kJ}}{\text{mol}}\right) + \left(\frac{25}{2} \text{ mol} \times \frac{495 \text{ kJ}}{\text{mol}}\right) -$$

$$\left[\left(16 \text{ mol} \times \frac{799 \text{ kJ}}{\text{mol}}\right) + \left(18 \text{ mol} \times \frac{467 \text{ kJ}}{\text{mol}}\right)\right]$$

$$= 7434 \text{ kJ} + 2429 \text{ kJ} + 6187.5 \text{ kJ} - (12\ 784 \text{ kJ} + 8406 \text{ kJ})$$

$$\Delta H = -5193.5 \text{ kJ}$$

**Step 4:** Calculate the enthalpy change, $\Delta H_{\text{per gram}}$, of the reaction for 1.00 g of octane.

$M_{C_8H_{18}(l)} = 114.26$ g/mol

$$\Delta H_{\text{per gram}} = \frac{\Delta H}{M_{C_8H_{18}(l)}}$$
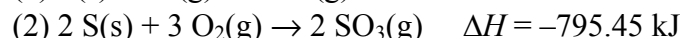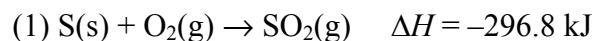
$$= \frac{-5193.5 \text{ kJ/mol}}{114.26 \text{ g/mol}}$$

$\Delta H = -45.5$ kJ/g

**Statement:** The enthalpy of combustion per gram of octane is –45.45 kJ. The combustion of 1.00 g of octane releases about twice the amount of energy released by the combustion of 1.00 g of ethanol. This means that octane is a more efficient fuel than ethanol.

**48. Solution:**

**Step 1:** Label the equations.

The desired equation is: $SO_2(g) + \dfrac{1}{2} O_2(g) \rightarrow SO_3(g)$

(1) $S(s) + O_2(g) \rightarrow SO_2(g)$    $\Delta H = -296.8$ kJ

(2) $2 S(s) + 3 O_2(g) \rightarrow 2 SO_3(g)$    $\Delta H = -795.45$ kJ

**Step 2:** Identify the equations that have reactants or products on the same side of the arrow as in the desired equation. Reverse (1) so that reactants are on the same side as in the desired equation. Reverse the sign of $\Delta H$.
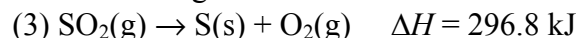
(3) $SO_2(g) \rightarrow S(s) + O_2(g)$    $\Delta H = 296.8$ kJ

**Step 3:** Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.

Equation (2) has 2 mol of $SO_3(g)$ on the product side, while the desired equation has 1 mol.

Multiply equation (2) and its $\Delta H$ by the factor $\dfrac{1}{2}$ :

(4) $S(s) + \dfrac{3}{2} O_2(g) \rightarrow SO_3(g)$    $\Delta H = -247.725$ kJ
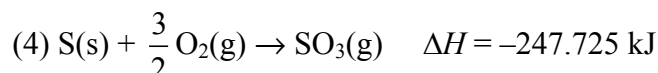
**Step 4:** Add equations (3) and (4), and their changes in enthalpies.

 (3) $SO_2(g)$          $\rightarrow \cancel{S(s) + O_2(g)}$    $\Delta H =$     296.8 kJ

(4) $\cancel{S(s)} + \cancel{\dfrac{3}{2}} O_2(g)$     $\rightarrow SO_3(g)$          $\Delta H = -397.725$ kJ
_____

  $SO_2(g) + \dfrac{1}{2} O_2(g) \rightarrow SO_3(g)$          $\Delta H = -100.925$ kJ
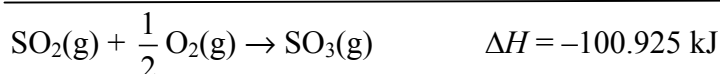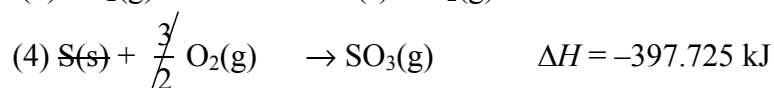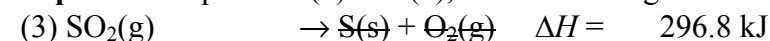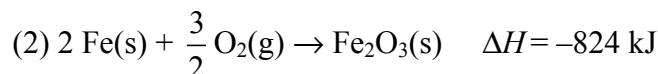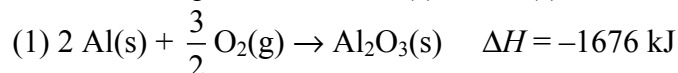
**Statement:** The enthalpy change for the required reaction is –100.9 kJ.
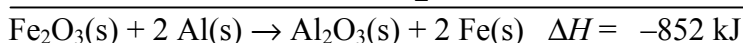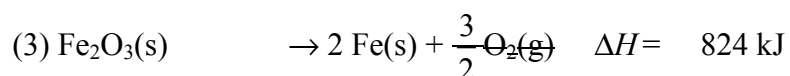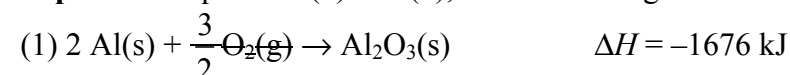
**49. Solution:**

**Step 1:** Label the equations.

The desired equation is: $Fe_2O_3(s) + 2 Al(s) \rightarrow Al_2O_3(s) + 2 Fe(s)$

(1) $2 Al(s) + \dfrac{3}{2} O_2(g) \rightarrow Al_2O_3(s)$    $\Delta H = -1676$ kJ

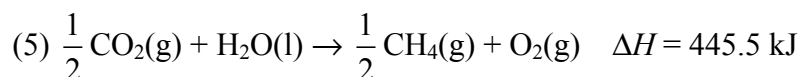(2) $2 Fe(s) + \dfrac{3}{2} O_2(g) \rightarrow Fe_2O_3(s)$    $\Delta H = -824$ kJ
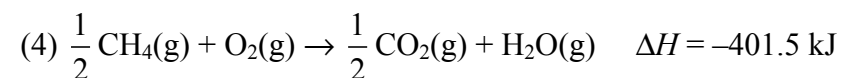
**Step 2:** Identify the equations that have reactants or products on the same side of the arrow as in the desired equation. Reverse (2) so that the reactant is on the same side as in the desired equation. Reverse the sign of $\Delta H$.
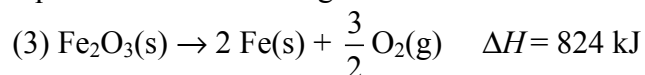
(3) $Fe_2O_3(s) \rightarrow 2\ Fe(s) + \dfrac{3}{2}\ O_2(g)$     $\Delta H = 824$ kJ

**Step 3:** Add equations (1) and (3), and their changes in enthalpies.

(1) $2\ Al(s) + \dfrac{3}{2}\ \cancel{O_2(g)} \rightarrow Al_2O_3(s)$ $\qquad\qquad$ $\Delta H = -1676$ kJ

(3) $Fe_2O_3(s) \qquad\qquad \rightarrow 2\ Fe(s) + \dfrac{3}{2}\ \cancel{O_2(g)}$   $\Delta H = \quad 824$ kJ

$\overline{\quad Fe_2O_3(s) + 2\ Al(s) \rightarrow Al_2O_3(s) + 2\ Fe(s) \quad \Delta H = \quad -852 \text{ kJ} \quad}$

**Statement:** The enthalpy change of the required reaction is –852 kJ.

**50. Solution:**

**Step 1:** The desired equation is: $H_2O(l) \rightarrow H_2O(g)$

Write the equation for the combustion of methane gas to liquid water.
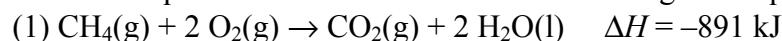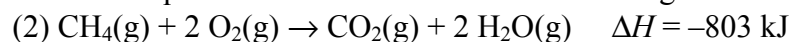
(1) $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$     $\Delta H = -891$ kJ

Write the equation for the combustion of methane gas to steam.

(2) $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$     $\Delta H = -803$ kJ

**Step 2:** Identify the equations that have reactants or products on the same side of the arrow as in the desired equation. Reverse (1) so that the reactant is on the same side as in the desired equation. Reverse the sign of $\Delta H$.

(3) $CO_2(g) + 2\ H_2O(l) \rightarrow CH_4(g) + 2\ O_2(g)$   $\Delta H = 891$ kJ
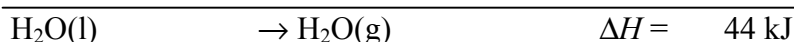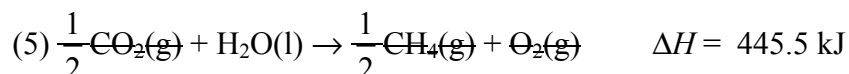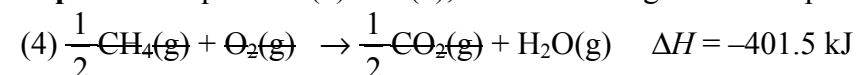
**Step 3:** Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.

Multiply equations (2) and (3) and their $\Delta H$ values by the factor $\dfrac{1}{2}$:

(4) $\dfrac{1}{2}\ CH_4(g) + O_2(g) \rightarrow \dfrac{1}{2}\ CO_2(g) + H_2O(g)$     $\Delta H = -401.5$ kJ

(5) $\dfrac{1}{2}\ CO_2(g) + H_2O(l) \rightarrow \dfrac{1}{2}\ CH_4(g) + O_2(g)$   $\Delta H = 445.5$ kJ

**Step 4:** Add equations (4) and (5), and their changes in enthalpies.

(4) $\dfrac{1}{2}\ \cancel{CH_4(g)} + \cancel{O_2(g)} \rightarrow \dfrac{1}{2}\ \cancel{CO_2(g)} + H_2O(g)$   $\Delta H = -401.5$ kJ

(5) $\dfrac{1}{2}\ \cancel{CO_2(g)} + H_2O(l) \rightarrow \dfrac{1}{2}\ \cancel{CH_4(g)} + \cancel{O_2(g)}$ $\qquad$ $\Delta H = \ 445.5$ kJ

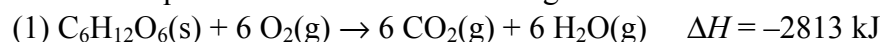$\overline{\quad H_2O(l) \qquad\qquad \rightarrow H_2O(g) \qquad\qquad \Delta H = \quad 44 \text{ kJ} \quad}$

**Statement:** The enthalpy of vaporization for water is 44 kJ/mol.

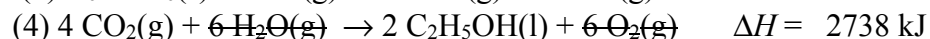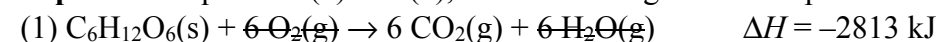**51. (a) Solution:**

**Step 1:** The desired equation is: $C_6H_{12}O_6(s) \rightarrow 2\ C_2H_5OH(l) + 2\ CO_2(g)$

Write the equation for the combustion of glucose.

(1) $C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(g)$     $\Delta H = -2813$ kJ

Write the equation for the combustion of ethanol.
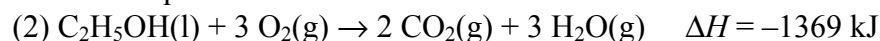
(2) $C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$     $\Delta H = -1369$ kJ

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**Step 2:** Identify the equations that have reactants or products on the same side of the arrow as in the desired equation. Reverse (2) so that the product is on the same side as in the desired equation. Reverse the sign of $\Delta H$.

(3) $2 CO_2(g) + 3 H_2O(g) \rightarrow C_2H_5OH(l) + 3 O_2(g)$     $\Delta H = 1369$ kJ

**Step 3:** Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.

Multiply equation (3) and its $\Delta H$ by the factor 2:

(4) $4 CO_2(g) + 6 H_2O(g) \rightarrow 2 C_2H_5OH(l) + 6 O_2(g)$     $\Delta H = 2738$ kJ
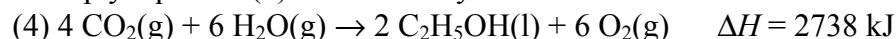
**Step 4:** Add equations (1) and (4), and their changes in enthalpies.

(1) $C_6H_{12}O_6(s) + \cancel{6 O_2(g)} \rightarrow 6 CO_2(g) + \cancel{6 H_2O(g)}$     $\Delta H = -2813$ kJ

(4) $\underline{4 CO_2(g) + \cancel{6 H_2O(g)} \rightarrow 2 C_2H_5OH(l) + \cancel{6 O_2(g)}}$     $\underline{\Delta H = \phantom{-}2738 \text{ kJ}}$

$\phantom{xx}C_6H_{12}O_6(s) \phantom{xxxxxx} \rightarrow 2 C_2H_5OH(l) + 2 CO_2(g)$     $\Delta H = \phantom{xx}-75$ kJ

**Statement:** The molar enthalpy of the fermentation reaction –75 kJ/mol.

**(b)** The reaction vessel would feel warm to the touch as the fermentation of glucose is an exothermic reaction. The reaction releases thermal energy.
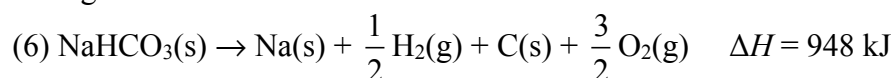
**52. Solution:**

**Step 1:** The desired equation is: (5) $2 NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$

Identify the equations that have reactants or products on the same side of the arrow as in the desired equation. Equations (2), (3), and (4) have the products on the same side of the arrow as in the desired equation, so use them as written.
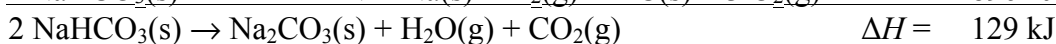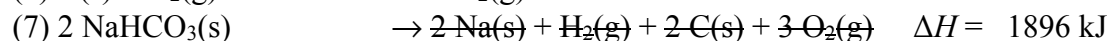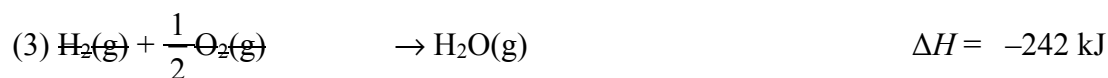
(2) $2 Na(s) + C(s) + \dfrac{3}{2} O_2(g) \rightarrow Na_2CO_3(s)$     $\Delta H = -1131$ kJ

(3) $H_2(g) + \dfrac{1}{2} O_2(g) \rightarrow H_2O(g)$     $\Delta H = -242$ kJ

(4) $C(s) + O_2(g) \rightarrow CO_2(g)$     $\Delta H = -394$ kJ
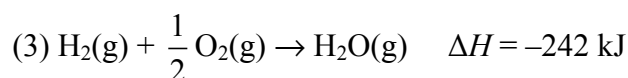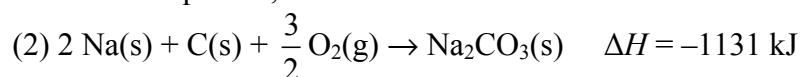
Reverse equation (1) so that the reactant is on the same side as in the desired equation. Reverse the sign of $\Delta H$.

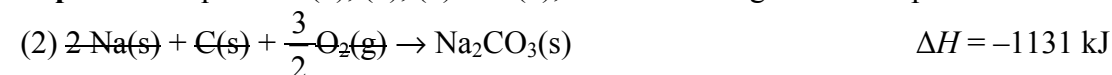(6) $NaHCO_3(s) \rightarrow Na(s) + \dfrac{1}{2} H_2(g) + C(s) + \dfrac{3}{2} O_2(g)$     $\Delta H = 948$ kJ

**Step 2:** Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.

Multiply equation (6) and its $\Delta H$ by the factor 2:

(7) $2 NaHCO_3(s) \rightarrow 2 Na(s) + H_2(g) + 2 C(s) + 3 O_2(g)$     $\Delta H = 1896$ kJ
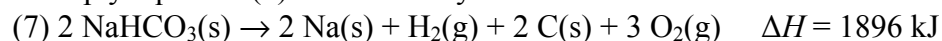
**Step 3:** Add equations (2), (3), (4) and (7), and their changes in enthalpies.

(2) $\cancel{2 Na(s)} + \cancel{C(s)} + \cancel{\dfrac{3}{2} O_2(g)} \rightarrow Na_2CO_3(s)$     $\Delta H = -1131$ kJ

(3) $\cancel{H_2(g)} + \cancel{\dfrac{1}{2} O_2(g)} \phantom{xxxx} \rightarrow H_2O(g)$     $\Delta H = \phantom{x}-242$ kJ

(4) $\cancel{C(s)} + \cancel{O_2(g)} \phantom{xxxxxx} \rightarrow CO_2(g)$     $\Delta H = \phantom{x}-394$ kJ

(7) $\underline{2 NaHCO_3(s) \phantom{xxxxxx} \rightarrow \cancel{2 Na(s)} + \cancel{H_2(g)} + \cancel{2 C(s)} + \cancel{3 O_2(g)}}$     $\underline{\Delta H = \phantom{x}1896 \text{ kJ}}$

$\phantom{xx}2 NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$     $\Delta H = \phantom{xx}129$ kJ

**Statement:** The enthalpy change for the required reaction is 129 kJ.
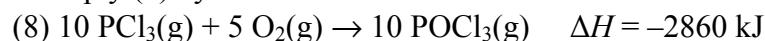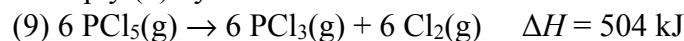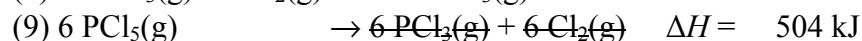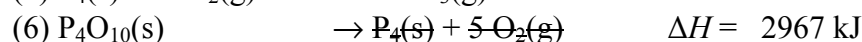
**53. Solution:**
**Step 1:** The desired equation is: (5) $P_4O_{10}(s) + 6\ PCl_5(g) \rightarrow 10\ POCl_3(g)$
Identify the equations that have reactants or products on the same side of the arrow as in the desired equation. Equation (4) has the product on the same side of the arrow as in the desired equation.

(4) $PCl_3(g) + \dfrac{1}{2} O_2(g) \rightarrow POCl_3(g)$ $\quad \Delta H = -286$ kJ
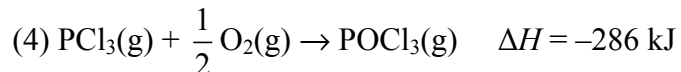
Reverse equations (2) and (3) so that the reactants are on the same side as in the desired equation. Reverse the sign of $\Delta H$.
(6) $P_4O_{10}(s) \rightarrow P_4(s) + 5\ O_2(g)$ $\quad \Delta H = 2967$ kJ
(7) $PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$ $\quad \Delta H = 84$ kJ
**Step 2:** Equation (1) has reactants and products that will be cancelled by adding other equations, so use it as written.
(1) $P_4(s) + 6\ Cl_2(g) \rightarrow 4\ PCl_3(g)$ $\quad \Delta H = -1226$ kJ
**Step 3:** Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.
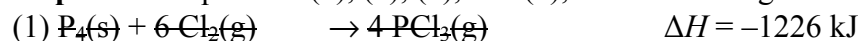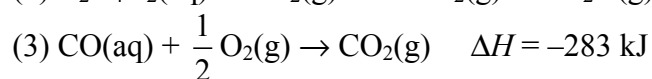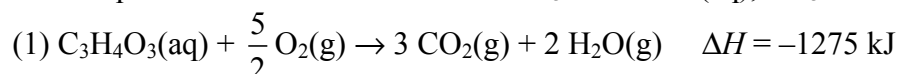Multiply (4) by the factor 10.
(8) $10\ PCl_3(g) + 5\ O_2(g) \rightarrow 10\ POCl_3(g)$ $\quad \Delta H = -2860$ kJ
Multiply (7) by the factor 6.
(9) $6\ PCl_5(g) \rightarrow 6\ PCl_3(g) + 6\ Cl_2(g)$ $\quad \Delta H = 504$ kJ
**Step 4:** Add equations (1), (6), (8), and (9), and their changes in enthalpies.
(1) $\cancel{P_4(s)} + \cancel{6\ Cl_2(g)}$ $\quad \rightarrow \cancel{4\ PCl_3(g)}$ $\qquad\qquad \Delta H = -1226$ kJ
(6) $P_4O_{10}(s)$ $\qquad\qquad \rightarrow \cancel{P_4(s)} + \cancel{5\ O_2(g)}$ $\qquad \Delta H = \ \ 2967$ kJ
(8) $\cancel{10\ PCl_3(g)} + \cancel{5\ O_2(g)} \rightarrow 10\ POCl_3(g)$ $\qquad\quad \Delta H = -2860$ kJ
(9) $\cancel{6\ PCl_5(g)}$ $\qquad\qquad \rightarrow \cancel{6\ PCl_3(g)} + \cancel{6\ Cl_2(g)}$ $\ \ \Delta H = \ \ \ \ \ 504$ kJ
$\qquad P_4O_{10}(s) + 6\ PCl_5(g) \rightarrow 10\ POCl_3(g)$ $\qquad\qquad \Delta H = \ \ -615$ kJ
**Statement:** The enthalpy change for the required reaction is –615 kJ.
**54. Solution:**
**Step 1:** The desired equation is: $CH_3COCOOH(aq) \rightarrow CH_3COOH(aq) + CO(aq)$
Write equations for the combustion of $CH_3COCOOH(aq)$, $CH_3COOH(aq)$, and $CO(aq)$.
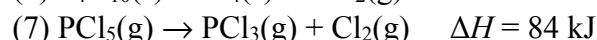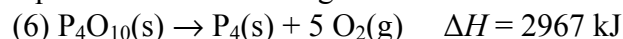
(1) $C_3H_4O_3(aq) + \dfrac{5}{2} O_2(g) \rightarrow 3\ CO_2(g) + 2\ H_2O(g)$ $\quad \Delta H = -1275$ kJ

(2) $C_2H_4O_2(aq) + 2\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(g)$ $\quad \Delta H = -875$ kJ

(3) $CO(aq) + \dfrac{1}{2} O_2(g) \rightarrow CO_2(g)$ $\quad \Delta H = -283$ kJ

**Step 2:** Identify the equations that have reactants or products on the same side of the arrow as in the desired equation. Equation (1) has the reactant on the same side of the arrow as in the desired equation. Reverse equations (2) and (3) so that the products are on the same side as in the desired equation. Reverse the sign of $\Delta H$.
(4) $2\ CO_2(g) + 2\ H_2O(g) \rightarrow C_2H_4O_2(aq) + 2\ O_2(g)$ $\quad \Delta H = 875$ kJ

(5) $CO_2(g) \rightarrow CO(aq) + \dfrac{1}{2} O_2(g)$ $\quad \Delta H = 283$ kJ
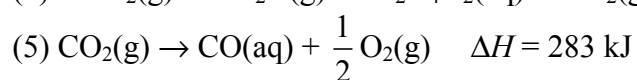
**Step 3:** Add equations (1), (4), and (5), and their changes in enthalpies.

(1) $C_3H_4O_3(aq) + \dfrac{5}{2} \cancel{O_2(g)} \rightarrow \cancel{3\ CO_2(g)} + \cancel{2\ H_2O(g)}$     $\Delta H = -1275$ kJ

(4) $\cancel{2\ CO_2(g)} + \cancel{2\ H_2O(g)} \rightarrow C_2H_4O_2(aq) + \cancel{2\ O_2(g)}$     $\Delta H =\ \ \ 875$ kJ

(5) $\cancel{CO_2(g)} \qquad\qquad\qquad \rightarrow CO(aq) + \dfrac{1}{2}\cancel{O_2(g)}$     $\Delta H =\ \ \ 283$ kJ
_____
$\phantom{(5)}\ C_3H_4O_3(aq) \qquad\qquad \rightarrow C_2H_4O_2(aq) + CO(aq)$     $\Delta H = -117$ kJ
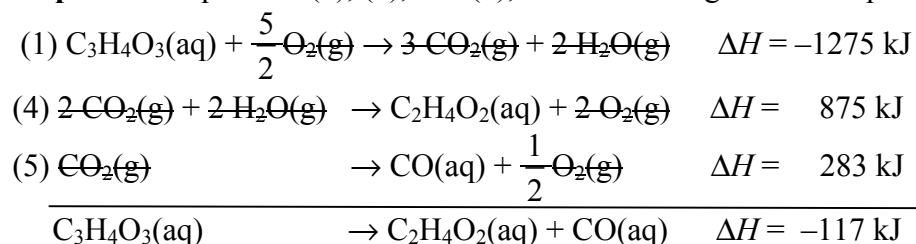
**Statement:** The enthalpy change for the required reaction is –117 kJ.

**55. (a)** Answers may vary. Students' flow charts should include:
For (i),
• Writing a balanced chemical equation that represents the reaction.
• Listing the standard enthalpies of formation of substances involved.
• The enthalpy equation for calculating enthalpy change using standard enthalpies of formation:

$\Delta H_r^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

For (ii),
• Determining the number of moles of reactants and products, the number of moles
of bonds broken or formed, and the molar bond energy for each.
• The formula for calculating the enthalpy change of a reaction using bond energies:

$\Delta H = \sum n \times D_{bonds\ broken} - \sum n \times D_{bonds\ formed}$

**(b)** Answers may vary. Students' answers should include the fact that bond energies are average
energies, so they are used to predict the approximate enthalpies while standard enthalpies of
formation are from results of actual experiments.

**56. (a) Given:** $\Delta H_f^\circ{}_{NH_3(g)} = -45.9$ kJ/mol; $\Delta H_f^\circ{}_{CH_4(g)} = -74.4$ kJ/mol;

$\Delta H_f^\circ{}_{HCN(g)} = 135.1$ kJ/mol; $\Delta H_f^\circ{}_{H_2O(g)} = -241.8$ kJ/mol; $\Delta H_f^\circ{}_{O_2(g)} = 0$ kJ/mol

**Required:** $\Delta H_r^\circ$

**Analysis:** $\Delta H_r^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

Since $O_2(g)$ is in its standard state, the equation can be written as:
$\Delta H_r^\circ = [n_{HCN(g)} \times \Delta H_f^\circ{}_{HCN(g)} + n_{H_2O(g)} \times \Delta H_f^\circ{}_{H_2O(g)}] -$

$\qquad [n_{NH_3(g)} \times \Delta H_f^\circ{}_{NH_3(g)} + n_{CH_4(g)} \times \Delta H_f^\circ{}_{CH_4(g)}]$

**Solution:**
Determine the amount of each reactant or product from the equation:
$2\ NH_3(g) + 3\ O_2(g) + CH_4(g) \rightarrow 2\ HCN(g) + 6\ H_2O(g)$
$\Delta H_r^\circ = [n_{HCN(g)} \times \Delta H_f^\circ{}_{HCN(g)} + n_{H_2O(g)} \times \Delta H_f^\circ{}_{H_2O(g)}] -$

$\qquad [n_{NH_3(g)} \times \Delta H_f^\circ{}_{NH_3(g)} + n_{CH_4(g)} \times \Delta H_f^\circ{}_{CH_4(g)}]$

$\qquad = [2(135.1\ \text{kJ}) + 6(-241.8\ \text{kJ})] - [2(-45.9\ \text{kJ}) + (-74.4\ \text{kJ})]$

$\qquad = 270.2\ \text{kJ} - 1450.8\ \text{kJ} - [-91.8\ \text{kJ} - 74.4\ \text{kJ}]$

$\Delta H_r^\circ = -1014.4$ kJ

**Statement:** $\Delta H^\circ$ for the required reaction is –1014.4 kJ.

**(b) Given:** $\Delta H_f°_{Ca_3(PO_4)_2(s)} = -4120.8$ kJ/mol; $\Delta H_f°_{H_2SO_4(l)} = -814.0$ kJ/mol;

$\Delta H_f°_{CaSO_4(s)} = -1434.1$ kJ/mol; $\Delta H_f°_{H_3PO_4(l)} = -1271.7$ kJ/mol

**Required:** $\Delta H_r°$

**Analysis:** $\Delta H_r° = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

**Solution:**

Determine the amount of each reactant or product from the equation:

$Ca_3(PO_4)_2(s) + 3\ H_2SO_4(l) \rightarrow 3\ CaSO_4(s) + 2\ H_3PO_4(l)$
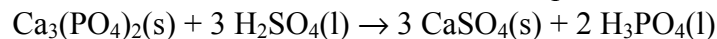
$\Delta H_r° = [n_{CaSO_4(s)} \times \Delta H_f°_{CaSO_4(s)} + n_{H_3PO_4(l)} \times \Delta H_f°_{H_3PO_4(l)}] -$

$\qquad [n_{Ca_3(PO_4)_2(s)} \times \Delta H_f°_{Ca_3(PO_4)_2(s)} + n_{H_2SO_4(l)} \times \Delta H_f°_{H_2SO_4(l)}]$

$\qquad = [3(-1434.1\text{ kJ}) + 2(-1271.7\text{ kJ})] - [(-4120.8\text{ kJ}) + 3(-814.0\text{ kJ})]$

$\qquad = -4302.3\text{ kJ} - 2543.4\text{ kJ} - [-4120.8\text{ kJ} - 2442.0\text{ kJ}]$

$\Delta H_r° = -282.9$ kJ

**Statement:** $\Delta H°$ for the required reaction is –282.9 kJ.

**(c) Given:** $\Delta H_f°_{NH_3(g)} = -45.9$ kJ/mol; $\Delta H_f°_{HCl(g)} = -92.3$ kJ/mol; $\Delta H_f°_{NH_4Cl(s)} = -314.4$ kJ/mol

**Required:** $\Delta H_r°$

**Analysis:** $\Delta H_r° = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

**Solution:**

Determine the amount of each reactant or product from the equation:

$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$
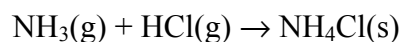
$\Delta H_r° = n_{NH_4Cl(s)} \times \Delta H_f°_{NH_4Cl(s)} - [n_{NH_3(g)} \times \Delta H_f°_{NH_3(g)} + n_{HCl(g)} \times \Delta H_f°_{HCl(g)}]$

$\qquad = -314.4\text{ kJ} - [(-45.9\text{ kJ}) + (-92.3\text{ kJ})]$

$\qquad = -314.4\text{ kJ} - [-45.9\text{ kJ} - 92.3\text{ kJ}]$

$\Delta H_r° = -176.2$ kJ

**Statement:** $\Delta H°$ for the required reaction is –176.2 kJ.

**(d) Given:** $\Delta H_f°_{C_2H_5OH(l)} = -235.2$ kJ/mol; $\Delta H_f°_{CO_2(g)} = -393.5$ kJ/mol;

$\Delta H_f°_{H_2O(l)} = -285.8$ kJ/mol; $\Delta H_f°_{O_2(g)} = 0$ kJ/mol

**Required:** $\Delta H_r°$

**Analysis:** $\Delta H_r° = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

Since $O_2(g)$ is in its standard state, the equation can be written as:

$\Delta H_r° = [n_{CO_2(g)} \times \Delta H_f°_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f°_{H_2O(l)}] - n_{C_2H_5OH(l)} \times \Delta H_f°_{C_2H_5OH(l)}$

**Solution:**
Determine the amount of each reactant or product from the equation:

$C_2H_5OH(l) + 2\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(l)$
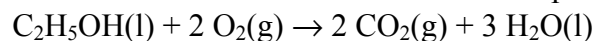
$\Delta H_r^\circ = [n_{CO_2(g)} \times \Delta H_f^{\circ}{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f^{\circ}{}_{H_2O(l)}] - n_{C_2H_5OH(l)} \times \Delta H_f^{\circ}{}_{C_2H_5OH(l)}$

$\quad\quad = [2(-393.5\ kJ) + 3(-285.8\ kJ)] - (-235.2\ kJ)$

$\quad\quad = -787.0\ kJ - 857.4\ kJ + 235.2\ kJ$

$\Delta H_r^\circ = -1409.2\ kJ$

**Statement:** $\Delta H^\circ$ for the required reaction is –1409.2 kJ.

**(e) Given:** $\Delta H_f^{\circ}{}_{SiCl_4(l)} = -687.0\ kJ/mol$; $\Delta H_f^{\circ}{}_{H_2O(l)} = -285.8\ kJ/mol$;

$\Delta H_f^{\circ}{}_{SiO_2(s)} = -910.7\ kJ/mol$; $\Delta H_f^{\circ}{}_{HCl(g)} = -92.3\ kJ/mol$

**Required:** $\Delta H_r^\circ$

**Analysis:** $\Delta H_r^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

**Solution:**
Determine the amount of each reactant or product from the equation:

$SiCl_4(l) + 2\ H_2O(l) \rightarrow SiO_2(s) + 4\ HCl(g)$
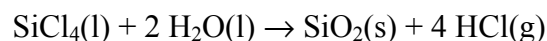
$\Delta H_r^\circ = [n_{SiO_2(s)} \times \Delta H_f^{\circ}{}_{SiO_2(s)} + n_{HCl(g)} \times \Delta H_f^{\circ}{}_{HCl(g)}] -$

$\quad\quad [n_{SiCl_4(l)} \times \Delta H_f^{\circ}{}_{SiCl_4(l)} + n_{H_2O(l)} \times \Delta H_f^{\circ}{}_{H_2O(l)}]$

$\quad\quad = [(-910.7\ kJ) + 4(-92.3\ kJ)] - [(-687.0\ kJ) + 2(-285.8\ kJ)]$

$\quad\quad = -910.7\ kJ - 369.2\ kJ - [-687.0\ kJ - 571.6\ kJ]$

$\Delta H_r^\circ = -21.3\ kJ$

**Statement:** $\Delta H^\circ$ for the required reaction is –21.3 kJ.

**(f) Given:** $\Delta H_f^{\circ}{}_{MgO(s)} = -601.6\ kJ/mol$; $\Delta H_f^{\circ}{}_{H_2O(l)} = -285.8\ kJ/mol$;

$\Delta H_f^{\circ}{}_{Mg(OH)_2(s)} = -924.5\ kJ/mol$

**Required:** $\Delta H_r^\circ$

**Analysis:** $\Delta H_r^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

**Solution:**
Determine the amount of each reactant or product from the equation:

$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(s)$
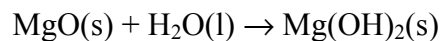
$\Delta H_r^\circ = n_{Mg(OH)_2(s)} \times \Delta H_f^{\circ}{}_{Mg(OH)_2(s)} - [n_{MgO(s)} \times \Delta H_f^{\circ}{}_{MgO(s)} + n_{H_2O(l)} \times \Delta H_f^{\circ}{}_{H_2O(l)}]$

$\quad\quad = -924.5\ kJ - [(-601.6\ kJ) + (-285.8\ kJ)]$

$\quad\quad = -924.5\ kJ - [-601.6\ kJ - 285.8\ kJ]$

$\Delta H_r^\circ = -37.1\ kJ$

**Statement:** $\Delta H^\circ$ for the required reaction is –37.1 kJ.

**57. (a) Given:** $\Delta H_f^\circ{}_{CO(g)} = -110.5$ kJ/mol; $\Delta H_f^\circ{}_{CH_3OH(l)} = -293.1$ kJ/mol; $\Delta H_f^\circ{}_{H_2(g)} = 0$ kJ/mol

**Required:** $\Delta H_r^\circ$

**Analysis:** $\Delta H_r^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

Since $H_2(g)$ is in its standard state, the equation can be written as:

$\Delta H_r^\circ = n_{CH_3OH(l)} \times \Delta H_f^\circ{}_{CH_3OH(l)} - n_{CO(g)} \times \Delta H_f^\circ{}_{CO(g)}$

**Solution:**

Determine the amount of each reactant or product from the equation:

$CO(g) + 2\ H_2(g) \rightarrow CH_3OH(l)$
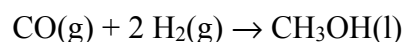
$\Delta H_r^\circ = n_{CH_3OH(l)} \times \Delta H_f^\circ{}_{CH_3OH(l)} - n_{CO(g)} \times \Delta H_f^\circ{}_{CO(g)}$

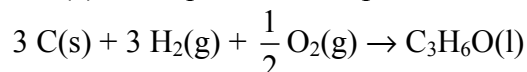$\qquad = -293.1\ \text{kJ} - (-110.5\ \text{kJ})$

$\Delta H_r^\circ = -128.6\ \text{kJ}$

**Statement:** $\Delta H^\circ$ for the required reaction is $-128.6$ kJ.

**(b)** Since $\Delta H^\circ$ for the required reaction is negative, the reaction is exothermic.

**58. (a)** The equation that represents the formation of acetone, $C_3H_6O(l)$, is:

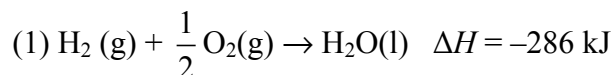$$3\ C(s) + 3\ H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow C_3H_6O(l)$$

**(b) Solution:**

**Step 1:** The desired equation is: $3\ C(s) + 3\ H_2(g) + \dfrac{1}{2}\ O_2(g) \rightarrow C_3H_6O(l)$

Identify the equations that have reactants or products on the same side of the arrow as in the desired equation.

Equations (1) and (2) have the reactants on the same side of the arrow as in the desired equation, so use them as written.
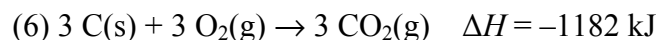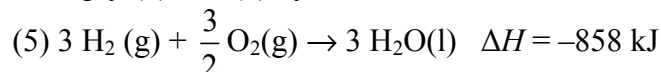
(1) $H_2(g) + \dfrac{1}{2}\ O_2(g) \rightarrow H_2O(l)\quad \Delta H = -286$ kJ

(2) $C(s) + O_2(g) \rightarrow CO_2(g)\quad \Delta H = -394$ kJ

Reverse equation (3) so that the product is on the same side as in the desired equation. Reverse the sign of $\Delta H$.

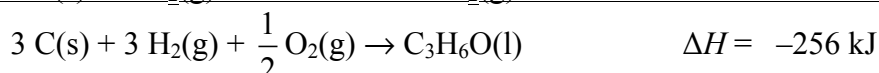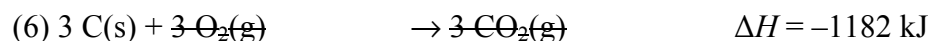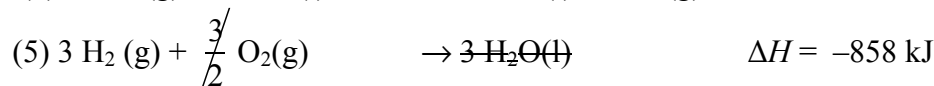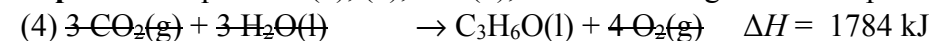(4) $3\ CO_2(g) + 3\ H_2O(l) \rightarrow C_3H_6O(l) + 4\ O_2(g)\quad \Delta H = 1784$ kJ

**Step 2:** Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.

Multiply (1) and (2) by the factor 3.

(5) $3\ H_2(g) + \dfrac{3}{2}\ O_2(g) \rightarrow 3\ H_2O(l)\quad \Delta H = -858$ kJ

(6) $3\ C(s) + 3\ O_2(g) \rightarrow 3\ CO_2(g)\quad \Delta H = -1182$ kJ

**Step 3:** Add equations (4), (5), and (6), and their changes in enthalpies.

(4) $\cancel{3\ CO_2(g)} + \cancel{3\ H_2O(l)} \qquad \rightarrow C_3H_6O(l) + \cancel{4\ O_2(g)}\quad \Delta H = 1784$ kJ

(5) $3\ H_2(g) + \cancel{\dfrac{3}{2}\ O_2(g)} \qquad \rightarrow \cancel{3\ H_2O(l)}\qquad\qquad \Delta H = -858$ kJ

(6) $3\ C(s) + \cancel{3\ O_2(g)} \qquad\qquad \rightarrow \cancel{3\ CO_2(g)}\qquad\qquad \Delta H = -1182$ kJ

$\qquad 3\ C(s) + 3\ H_2(g) + \dfrac{1}{2}\ O_2(g) \rightarrow C_3H_6O(l)\qquad\qquad \Delta H = -256$ kJ

**Statement:** The standard enthalpy of formation of acetone is –256 kJ/mol.

**59. Given:** $\Delta H_f{}^\circ{}_{CO_2(g)} = -393.5$ kJ/mol; $\Delta H_f{}^\circ{}_{H_2O(l)} = -285.8$ kJ/mol; $\Delta H_c{}^\circ = -3224$ kJ/mol;

$\Delta H_f{}^\circ{}_{O_2(g)} = 0$ kJ/mol

**Required:** $\Delta H_f{}^\circ{}_{C_6H_5COOH(s)}$

**Analysis:** $\Delta H_r{}^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

Since $O_2(g)$ is in its standard state, the equation can be written as:

$\Delta H_c{}^\circ = [n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^\circ{}_{H_2O(l)}] - n_{C_6H_5COOH(s)} \times \Delta H_f{}^\circ{}_{C_6H_5COOH(s)}$

**Solution:**

**Step 1:** Write a balanced chemical equation for the combustion of solid benzoic acid so that benzoic acid has a coefficient of 1.
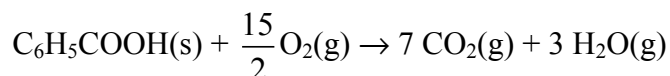
$$C_6H_5COOH(s) + \frac{15}{2} O_2(g) \rightarrow 7\ CO_2(g) + 3\ H_2O(g)$$

**Step 2:** Substitute the appropriate values for standard enthalpy into the equation and solve.

$$\Delta H_c{}^\circ = [n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^\circ{}_{H_2O(l)}] -$$

$$n_{C_6H_5COOH(s)} \times \Delta H_f{}^\circ{}_{C_6H_5COOH(s)}$$

$$-3224\text{ kJ} = [7(-393.5\text{ kJ}) + 3(-241.8\text{ kJ})] - \Delta H_f{}^\circ{}_{C_6H_5COOH(s)}$$

$$-3224\text{ kJ} = [-2754.5\text{ kJ} - 725.4\text{ kJ}] - 1\text{ mol} \times \Delta H_f{}^\circ{}_{C_6H_5COOH(s)}$$

$$1\text{ mol} \times \Delta H_f{}^\circ{}_{C_6H_5COOH(s)} = -3479.9\text{ kJ} + 3224\text{ kJ}$$

$$\Delta H_f{}^\circ{}_{C_6H_5COOH(s)} = -256\text{ kJ/mol}$$

**Statement:** $\Delta H_f{}^\circ$ for solid benzoic acid is –256 kJ/mol.

**60. (a) Given:** $\Delta H_f{}^\circ{}_{CO_2(g)} = -393.5$ kJ/mol; $\Delta H_f{}^\circ{}_{H_2O(l)} = -285.8$ kJ/mol;

$\Delta H_f{}^\circ{}_{C_6H_{12}O_6(s)} = -1273.1$ kJ/mol; $\Delta H_f{}^\circ{}_{O_2(g)} = 0$ kJ/mol

**Required:** $\Delta H_r{}^\circ$

**Analysis:** $\Delta H_r{}^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

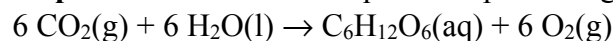Since $O_2(g)$ is in its standard state, the equation can be written as:

$\Delta H_c{}^\circ = [n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^\circ{}_{H_2O(l)}] - n_{C_6H_{12}O_6(aq)} \times \Delta H_f{}^\circ{}_{C_6H_{12}O_6(aq)}$

Assume that $\Delta H_f{}^\circ{}_{C_6H_{12}O_6(aq)} = \Delta H_f{}^\circ{}_{C_6H_{12}O_6(s)} = -1273.1$ kJ/mol.

**Solution:**

**Step 1:** Write a balanced equation representing the photosynthesis reaction:

$6\ CO_2(g) + 6\ H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g)$

**Step 2:** Determine the amount of each reactant or product from the equation.

$$\Delta H_r{}^\circ = n_{C_6H_{12}O_6(aq)} \times \Delta H_f{}^\circ{}_{C_6H_{12}O_6(aq)} - [n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^\circ{}_{H_2O(l)}]$$

$$= (-1273.1\text{ kJ}) - [6(-393.5\text{ kJ}) + 6(-285.8\text{ kJ})]$$

$$= -1273.1\text{ kJ} - [-2361.0\text{ kJ} - 1714.8\text{ kJ}]$$

$$\Delta H_r{}^\circ = 2802.7\text{ kJ}$$

**Statement:** $\Delta H^\circ$ for photosynthesis using $\Delta H_f{}^\circ$ values is 2802.7 kJ.

**(b) Solution:**
**Step 1:** Use the balanced chemical equation representing the photosynthesis reaction from part (a).
$6 CO_2(g) + 6 H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6 O_2(g)$
Determine the bonding of each substance.
**Step 2:** Determine the number of moles of reactants and products, the number of moles of bonds broken or formed, and the molar bond energy for each. Organize this information in a table.
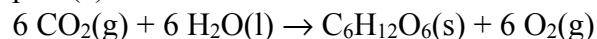
| | Substance | Number of bonds per mole ($n_{substance}$) | Amount of bonds in reaction | Bond energy per mole |
|---|---|---|---|---|
| reactants | $CO_2$ | 2 mol C=O bonds | 12 mol | 799 kJ/mol |
| | $H_2O$ | 2 mol H–O bonds | 12 mol | 467 kJ/mol |
| products | $C_6H_{12}O_6$ | 7 mol C–H bonds | 7 mol | 413 kJ/mol |
| | | 5 mol C–C bonds | 5 mol | 347 kJ/mol |
| | | 5 mol H–O bonds | 5 mol | 467 kJ/mol |
| | | 7 mol C–O bonds | 7 mol | 358 kJ/mol |
| | $O_2$ | 1 mol O=O bonds | 6 mol | 495 kJ/mol |

**Step 3:** Calculate the enthalpy change, $\Delta H$, of the reaction.
$$\Delta H = \sum n \times D_{\text{bonds broken}} - \sum n \times D_{\text{bonds formed}}$$
$$= 12 \text{ mol} \times D_{C=O} + 12 \text{ mol} \times D_{H-O} -$$
$$(7 \text{ mol} \times D_{C-H} + 5 \text{ mol} \times D_{C-C} + 5 \text{ mol} \times D_{H-O} + 7 \text{ mol} \times D_{C-O} + 6 \text{ mol} \times D_{O=O})$$
$$= \left(12 \text{ mol} \times \frac{799 \text{ kJ}}{\text{mol}}\right) + \left(12 \text{ mol} \times \frac{467 \text{ kJ}}{\text{mol}}\right) - \left[\left(7 \text{ mol} \times \frac{413 \text{ kJ}}{\text{mol}}\right) + \left(5 \text{ mol} \times \frac{347 \text{ kJ}}{\text{mol}}\right) + \right.$$
$$\left.\left(5 \text{ mol} \times \frac{467 \text{ kJ}}{\text{mol}}\right) + \left(7 \text{ mol} \times \frac{358 \text{ kJ}}{\text{mol}}\right) + \left(6 \text{ mol} \times \frac{495 \text{ kJ}}{\text{mol}}\right)\right]$$
$$= 9588 \text{ kJ} + 5604 \text{ kJ} - (2891 \text{ kJ} + 1735 \text{ kJ} + 2335 \text{ kJ} + 2506 \text{ kJ} + 2970 \text{ kJ})$$
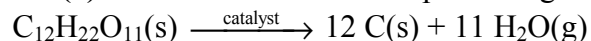$\Delta H = 2755 \text{ kJ}$
**Statement:** $\Delta H°$ for photosynthesis using bond energies is 2755 kJ.
**(c)** Answers may vary. Sample answer: The enthalpy determined using bond energies is close to the enthalpy determined using enthalpies of formation. The discrepancy between the two methods is due to the fact that bond energies are reported as average bond energies. The exact bond energy of a given bond depends on the types of atoms and bonds in the same molecule.
**(d)** Since the $\Delta H°$ value for photosynthesis is positive, the reaction is endothermic.
**61. (a)** The chemical reaction representing the process is:
$C_{12}H_{22}O_{11}(s) \xrightarrow{\text{catalyst}} 12 C(s) + 11 H_2O(g)$
**(b) Given:** $\Delta H_f°{}_{C_{12}H_{22}O_{11}(s)} = -2225.5$ kJ/mol; $\Delta H_f°{}_{H_2O(g)} = -241.8$ kJ/mol; $\Delta H_f°{}_{C(s)} = 0$ kJ/mol
**Required:** $\Delta H_r°$
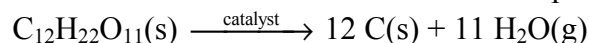**Analysis:** $\Delta H_r° = \sum n_{\text{products}} \times \Delta H_{\text{products}} - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}$
Since C(s) is in its standard state, the equation can be written as:
$\Delta H_c° = n_{H_2O(g)} \times \Delta H_f°{}_{H_2O(g)} - n_{C_{12}H_{22}O_{11}(s)} \times \Delta H_f°{}_{C_{12}H_{22}O_{11}(s)}$

**Solution:**
Determine the amount of each reactant or product from the equation:

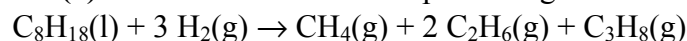$$C_{12}H_{22}O_{11}(s) \xrightarrow{\text{catalyst}} 12\ C(s) + 11\ H_2O(g)$$

$$\Delta H_r^\circ = n_{H_2O(g)} \times \Delta H_f^\circ{}_{H_2O(g)} - n_{C_{12}H_{22}O_{11}(s)} \times \Delta H_f^\circ{}_{C_{12}H_{22}O_{11}(s)}$$

$$= 11(-241.8\ \text{kJ}) - (-2225.5\ \text{kJ})$$

$$= -2659.8\ \text{kJ} + 2225.5\ \text{kJ}$$

$$\Delta H_r^\circ = -434.3\ \text{kJ}$$

**Statement:** The molar enthalpy change for the reaction is –434.3 kJ.

**62. (a)** The chemical reaction representing the cracking process of gasoline is:

$$C_8H_{18}(l) + 3\ H_2(g) \rightarrow CH_4(g) + 2\ C_2H_6(g) + C_3H_8(g)$$

**(b) Given:** $\Delta H_f^\circ{}_{C_8H_{18}(l)} = -250.1\ \text{kJ/mol}$; $\Delta H_f^\circ{}_{CH_4(g)} = -74.4\ \text{kJ/mol}$;

$\Delta H_f^\circ{}_{C_2H_6(g)} = -83.8\ \text{kJ/mol}$; $\Delta H_f^\circ{}_{C_3H_8(g)} = -104.7\ \text{kJ/mol}$; $\Delta H_f^\circ{}_{H_2(g)} = 0\ \text{kJ/mol}$

**Required:** $\Delta H_r^\circ$

**Analysis:** $\Delta H_r^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

Since $H_2(g)$ is in its standard state, the equation can be written as:

$$\Delta H_r^\circ = [n_{CH_4(g)} \times \Delta H_f^\circ{}_{CH_4(g)} + n_{C_2H_6(g)} \times \Delta H_f^\circ{}_{C_2H_6(g)} + n_{C_3H_8(g)} \times \Delta H_f^\circ{}_{C_3H_8(g)}] -$$

$$n_{C_8H_{18}(l)} \times \Delta H_f^\circ{}_{C_8H_{18}(l)}$$

**Solution:**
Determine the amount of each reactant or product from the equation:

$$C_8H_{18}(l) + 3\ H_2(g) \rightarrow CH_4(g) + 2\ C_2H_6(g) + C_3H_8(g)$$

$$\Delta H_r^\circ = [n_{CH_4(g)} \times \Delta H_f^\circ{}_{CH_4(g)} + n_{C_2H_6(g)} \times \Delta H_f^\circ{}_{C_2H_6(g)} + n_{C_3H_8(g)} \times \Delta H_f^\circ{}_{C_3H_8(g)}] -$$

$$n_{C_8H_{18}(l)} \times \Delta H_f^\circ{}_{C_8H_{18}(l)}$$

$$= [(-74.4\ \text{kJ}) + 2(-83.8\ \text{kJ}) + (-104.7\ \text{kJ})] - (-250.1\ \text{kJ})$$

$$= -74.4\ \text{kJ} - 167.6\ \text{kJ} - 104.7\ \text{kJ} + 250.1\ \text{kJ}$$

$$\Delta H_r^\circ = -96.6\ \text{kJ}$$

**Statement:** The molar enthalpy change that occurs during the cracking process is –96.6 kJ.

**63. Given:** $\Delta H_f^\circ{}_{H_2O(l)} = -285.8\ \text{kJ/mol}$; $\Delta H_f^\circ{}_{CH_4(g)} = -74.4\ \text{kJ/mol}$; $\Delta H_f^\circ{}_{O_2(g)} = 0\ \text{kJ/mol}$

**Required:** $\Delta H_c^\circ{}_{hydrogen}$; $\Delta H_c^\circ{}_{methane}$

**Analysis:** $\Delta H_r^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

Since $O_2(g)$ is in its standard state, the equations can be written as:

$$\Delta H_c^\circ{}_{hydrogen} = n_{H_2O(l)} \times \Delta H_f^\circ{}_{H_2O(l)}$$

$$\Delta H_c^\circ{}_{methane} = [n_{CO_2(g)} \times \Delta H_f^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f^\circ{}_{H_2O(l)}] - n_{methane} \times \Delta H_f^\circ{}_{methane}$$

**Solution:**
**Step 1:** Write a balanced chemical equation for the combustion of each substance so that the substance has a coefficient of 1.

$$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$$

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$$
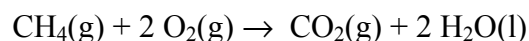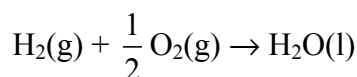
**Step 2:** Substitute the appropriate values for standard enthalpy into the equation for each substance and solve.

Hydrogen:

$$\Delta H_{c}^{\circ}{}_{\text{hydrogen}} = n_{\text{H}_2\text{O(l)}} \times \Delta H_{f}^{\circ}{}_{\text{H}_2\text{O(l)}}$$

$$\Delta H_{c}^{\circ}{}_{\text{hydrogen}} = -285.8 \text{ kJ}$$

Methane:

$$\Delta H_{c}^{\circ}{}_{\text{methane}} = [n_{\text{CO}_2\text{(g)}} \times \Delta H_{f}^{\circ}{}_{\text{CO}_2\text{(g)}} + n_{\text{H}_2\text{O(l)}} \times \Delta H_{f}^{\circ}{}_{\text{H}_2\text{O(l)}}] - n_{\text{methane}} \times \Delta H_{f}^{\circ}{}_{\text{methane}}$$

$$= [(-393.5 \text{ kJ}) + 2(-285.8 \text{ kJ})] - (-74.4 \text{ kJ})$$

$$= -393.5 \text{ kJ} - 571.6 \text{ kJ} + 74.4 \text{ kJ}$$

$$\Delta H_{c}^{\circ}{}_{\text{methane}} = -890.7 \text{ kJ}$$

**Step 3:** Convert to enthalpy per gram of each substance.

$$M_{\text{H}_2\text{(g)}} = 2.02 \text{ g/mol}$$

$$\Delta H_{c}^{\circ}{}_{\text{per gram H}_2\text{(g)}} = \frac{\Delta H_{c}^{\circ}{}_{\text{per mole H}_2\text{(g)}}}{M_{\text{H}_2\text{(g)}}}$$

$$= \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{2.02 \text{ g}}$$

$$\Delta H_{c}^{\circ}{}_{\text{per gram H}_2\text{(g)}} = -141.5 \text{ kJ/g}$$

$$M_{\text{CH}_4\text{(g)}} = 16.05 \text{ g/mol}$$

$$\Delta H_{c}^{\circ}{}_{\text{per gram CH}_4\text{(g)}} = \frac{\Delta H_{c}^{\circ}{}_{\text{per mole CH}_4\text{(g)}}}{M_{\text{CH}_4\text{(g)}}}$$

$$= \frac{-890.7 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{16.05 \text{ g}}$$

$$\Delta H_{c}^{\circ}{}_{\text{per gram CH}_4\text{(g)}} = -55.50 \text{ kJ/g}$$

**Step 4:** Calculate the ratio of energy released by the combustion of 1.00 g of hydrogen to the energy released by the combustion of 1.00 g of methane.

$$\frac{-141.5 \text{ kJ/g}}{-55.50 \text{ kJ/g}} \approx 2.5$$

**Statement:** Therefore, the combustion reaction of hydrogen gas produces approximately 2.5 times more energy per gram than the combustion of methane gas.

**64. Given:** $\Delta H_c^\circ{}_{C_3H_8(g)} = -2221 \text{ kJ/mol}$; energy output $= 8.7 \times 10^4 \text{ kJ}$; efficiency $= 60 \text{ %}$

**Required:** $m_{C_3H_8(g)}$

**Analysis:** $\dfrac{\text{energy output}}{\text{energy input}} = \text{efficiency}$

**Solution:**

**Step 1:** Calculate the energy input from the combustion of propane.

$\dfrac{\text{energy output}}{\text{energy input}} = 60 \text{ %}$

$\text{energy input} = \dfrac{\text{energy output}}{60 \text{ %}}$

$\qquad\qquad\quad = \dfrac{8.7 \times 10^4 \text{ kJ}}{0.60}$

$\text{energy input} = 14.5 \times 10^4 \text{ kJ (1 extra digit carried)}$

**Step 2:** Calculate the amount of propane that would produce this amount of energy, $n_{C_3H_8(g)}$.

$n_{C_3H_8(g)} = \dfrac{14.5 \times 10^4 \text{ kJ}}{-2221 \text{ kJ/mol}}$

$n_{C_3H_8(g)} = 65.29 \text{ mol (2 extra digits carried)}$

**Step 3:** Calculate the mass of propane required, $m_{C_3H_8(g)}$.

$M_{C_3H_8(g)} = 44.11 \text{ g/mol}$

$m_{C_3H_8(g)} = n_{C_3H_8(g)} M_{C_3H_8(g)}$

$\qquad\quad = (65.29 \text{ mol}) \left( \dfrac{44.11 \text{ g}}{1 \text{ mol}} \right)$

$m_{C_3H_8(g)} = 2900 \text{ g}$

**Statement:** The mass of propane that must be burned to furnish $8.7 \times 10^4$ kJ of energy is 2900 g.

**65.** Answers may vary. Students' concept maps should include environmental consequences of burning fossil fuels such as these:
• Fossil fuels are non-renewable.
• Sulfur dioxide gas comes from the combustion of fossil fuels containing sulfur impurities. Once in the atmosphere, sulfur dioxide undergoes further chemical reactions to form acid precipitation, which can cause significant damage to aquatic ecosystems.
• The combustion of all fossil fuels produces carbon dioxide gas and water vapour, which contribute to global warming and climate change.

**66. (a) Given:** $\Delta H_{fus} = 1.7 \times 10^9$ kJ/mol; $m_{He(g)} = 500$ g

**Required:** $\Delta H_{He(g)}$

**Analysis:** $\Delta H = n\Delta H_{fus}$

**Solution:**

**Step 1:** Calculate the amount of He(g) in 500 g.

$M_{He(g)} = 4.00$ g/mol

$$n_{He(g)} = \frac{m_{He(g)}}{M_{He(g)}}$$

$$n_{He(g)} = \frac{500 \ \cancel{g}}{4.00 \ \cancel{g}/mol}$$

$$n_{He(g)} = 125 \ mol$$

**Step 2:** Calculate the energy released by 125 mol of He(g), $\Delta H_{He(g)}$.

$$\Delta H_{He(g)} = n_{He(g)}\Delta H_{fus}$$

$$= (125 \ \cancel{mol}) \left( 1.7 \times 10^9 \ \frac{kJ}{\cancel{mol}} \right)$$

$$\Delta H_{He(g)} = 2.1 \times 10^{11} \ kJ$$

**Statement:** The energy released when 500 g of helium atoms are produced by nuclear fusion is $2.1 \times 10^{11}$ kJ.

**(b)** Nuclear fusion releases much more energy than any known fuel sources.

**(c)** Fusion reaction is not used to generate electricity because the intense pressure and heat needed for fusion to occur have made it difficult to develop fusion reactors to generate energy.

**Evaluation**

**67. (a) Given:** $\Delta H_f^\circ{}_{C_3H_8(g)} = -104.7$ kJ/mol; $\Delta H_f^\circ{}_{C_8H_{18}(l)} = -250.1$ kJ/mol;

$\Delta H_f^\circ{}_{CO_2(g)} = -393.5$ kJ/mol; $\Delta H_f^\circ{}_{H_2O(l)} = -285.8$ kJ/mol; $\Delta H_f^\circ{}_{O_2(g)} = 0$ kJ/mol

**Required:** $\Delta H_c^\circ{}_{per \ gram \ C_3H_8(g)}$; $\Delta H_c^\circ{}_{per \ gram \ C_8H_{18}(l)}$

**Analysis:** $\Delta H_r^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$

Since O$_2$(g) is in its standard state, the equations can be written as:

$\Delta H_c^\circ{}_{C_3H_8(g)} = [n_{CO_2(g)} \times \Delta H_f^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f^\circ{}_{H_2O(l)}] - n_{C_3H_8(g)} \times \Delta H_f^\circ{}_{C_3H_8(g)}$

$\Delta H_c^\circ{}_{C_8H_{18}(l)} = [n_{CO_2(g)} \times \Delta H_f^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f^\circ{}_{H_2O(l)}] - n_{C_8H_{18}(l)} \times \Delta H_f^\circ{}_{C_8H_{18}(l)}$

**Solution:**

**Step 1:** Write a balanced chemical equation for the combustion of each substance so that the substance has a coefficient of 1.

C$_3$H$_8$(g) + 5 O$_2$(g) → 3 CO$_2$(g) + 4 H$_2$O(l)

C$_8$H$_{18}$(l) + $\dfrac{25}{2}$ O$_2$(g) → 8 CO$_2$(g) + 9 H$_2$O(g)
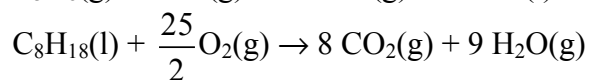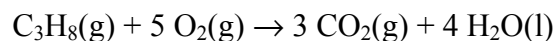
**Step 2:** Substitute the appropriate values for standard enthalpy of formation into the equation and solve.

Propane:

$$\Delta H_c{}^\circ{}_{C_3H_8(g)} = [n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^\circ{}_{H_2O(l)}] - n_{C_3H_8(g)} \times \Delta H_f{}^\circ{}_{C_3H_8(g)}$$

$$= [3(-393.5 \text{ kJ}) + 4(-285.8 \text{ kJ})] - (-104.7 \text{ kJ})$$

$$= -1180.5 \text{ kJ} - 1143.2 \text{ kJ} + 104.7 \text{ kJ}$$

$$\Delta H_c{}^\circ{}_{C_3H_8(g)} = -2219.0 \text{ kJ}$$

Gasoline:

$$\Delta H_c{}^\circ{}_{C_8H_{18}(l)} = [n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^\circ{}_{H_2O(l)}] - n_{C_8H_{18}(l)} \times \Delta H_f{}^\circ{}_{C_8H_{18}(l)}$$

$$= [8(-393.5 \text{ kJ}) + 9(-285.8 \text{ kJ})] - (-250.1 \text{ kJ})$$

$$= -3148.0 \text{ kJ} - 2572.2 \text{ kJ} + 250.1 \text{ kJ}$$

$$\Delta H_c{}^\circ{}_{C_8H_{18}(l)} = -5470.1 \text{ kJ}$$

**Step 3:** Convert to enthalpy per gram of each substance.

Propane:

$$M_{C_3H_8(g)} = 44.11 \text{ g/mol}$$

$$\Delta H_c{}^\circ{}_{\text{per gram } C_3H_8(g)} = \frac{\Delta H_c{}^\circ{}_{\text{per mole } C_3H_8(g)}}{M_{C_3H_8(g)}}$$

$$= \frac{-2219.0 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{44.11 \text{ g}}$$

$$\Delta H_c{}^\circ{}_{\text{per gram } C_3H_8(g)} = -50.3 \text{ kJ/g}$$

Gasoline:

$$M_{C_8H_{18}(l)} = 114.26 \text{ g/mol}$$

$$\Delta H_c{}^\circ{}_{\text{per gram } C_8H_{18}(l)} = \frac{\Delta H_c{}^\circ{}_{\text{per mole } C_8H_{18}(l)}}{M_{C_8H_{18}(l)}}$$

$$= \frac{-5470.1 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{114.26 \text{ g}}$$

$$\Delta H_c{}^\circ{}_{\text{per gram } C_8H_{18}(l)} = -47.9 \text{ kJ/g}$$

**Statement:** The amount of energy per gram released by complete combustion of propane is −50.3 kJ and that of gasoline is –47.9 kJ.

**(b)** The more energy-efficient fuel is propane, as it releases more energy per gram than gasoline.

**(c)** Answers may vary. Sample answer: I do not think gasoline should be replaced by propane gas in the trucking industry because it is more dangerous to transport a canister of compressed propane gas in a car than a tank of liquid gasoline, as compressed gas fuel is much more explosive than liquid fuel.

**68. (a) Given:** $\Delta H_f{}^\circ{}_{CH_4(g)} = -74.4$ kJ/mol; $\Delta H_f{}^\circ{}_{CO(g)} = -110.5$ kJ/mol;

$\Delta H_f{}^\circ{}_{CO_2(g)} = -393.5$ kJ/mol; $\Delta H_f{}^\circ{}_{H_2O(l)} = -285.8$ kJ/mol; $\Delta H_f{}^\circ{}_{O_2(g)} = 0$ kJ/mol

**Required:** $\Delta H_c{}^\circ{}_{CH_4(g)}$; $\Delta H_c{}^\circ{}_{CO(g)}$

**Analysis:** $\Delta H_r{}^\circ = \sum n_{products} \times \Delta H_{products} - \sum n_{reactants} \times \Delta H_{reactants}$
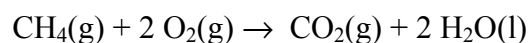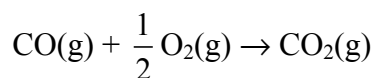
Since $O_2(g)$ is in its standard states, the equations can be written as:

$\Delta H_c{}^\circ{}_{CH_4(g)} = [n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^\circ{}_{H_2O(l)}] - n_{CH_4(g)} \times \Delta H_f{}^\circ{}_{CH_4(g)}$

$\Delta H_c{}^\circ{}_{CO(g)} = n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} - n_{CO(g)} \times \Delta H_f{}^\circ{}_{CO(g)}$

**Solution:**

**Step 1:** Write a balanced chemical equation for the combustion of each substance so that the substance has a coefficient of 1.

$CO(g) + \dfrac{1}{2} O_2(g) \rightarrow CO_2(g)$

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$

**Step 2:** Substitute the appropriate values for standard enthalpy into the equation for each substance and solve.

Methane:

$\Delta H_c{}^\circ{}_{CH_4(g)} = [n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} + n_{H_2O(l)} \times \Delta H_f{}^\circ{}_{H_2O(l)}] - n_{CH_4(g)} \times \Delta H_f{}^\circ{}_{CH_4(g)}$

$= [(-393.5\text{ kJ}) + 2(-285.8\text{ kJ})] - (-74.4\text{ kJ})$

$= -393.5\text{ kJ} - 571.6\text{ kJ} + 74.4\text{ kJ}$

$\Delta H_c{}^\circ{}_{CH_4(g)} = -890.7$ kJ

Carbon monoxide:

$\Delta H_c{}^\circ{}_{CO(g)} = n_{CO_2(g)} \times \Delta H_f{}^\circ{}_{CO_2(g)} - n_{CO(g)} \times \Delta H_f{}^\circ{}_{CO(g)}$

$= (-393.5\text{ kJ}) - (-110.5\text{ kJ})$

$= -393.5\text{ kJ} + 110.5\text{ kJ}$

$\Delta H_c{}^\circ{}_{CO(g)} = -283.0$ kJ

**Statement:** The enthalpy change for the combustion of methane is –890.7 kJ/mol and that of carbon monoxide is –283.0 kJ/mol.

**(b) Given:** $\Delta H_c{}^\circ{}_{CH_4(g)} = -890.7$ kJ/mol; $\Delta H_c{}^\circ{}_{CO(g)} = -283.0$ kJ/mol

**Required:** $\Delta H_c{}^\circ{}_{\text{per gram } CH_4(g)}$; $\Delta H_c{}^\circ{}_{\text{per gram } CO(g)}$

**Analysis:** $\Delta H_c{}^\circ{}_{\text{per gram}} = \dfrac{\Delta H_c{}^\circ{}_{\text{per mole}}}{M}$

**Solution:**

$M_{CH_4(g)} = 16.05$ g/mol

$$\Delta H_c{}^\circ{}_{\text{per gram } CH_4(g)} = \frac{\Delta H_c{}^\circ{}_{\text{per mole } CH_4(g)}}{M_{CH_4(g)}}$$

$$= \frac{-890.7 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{16.05 \text{ g}}$$

$\Delta H_c{}^\circ{}_{\text{per gram } CH_4(g)} = -55.50$ kJ/g

$M_{CO(g)} = 28.01$ g/mol

$$\Delta H_c{}^\circ{}_{\text{per gram } CO(g)} = \frac{\Delta H_c{}^\circ{}_{\text{per mole } CO(g)}}{M_{CO(g)}}$$

$$= \frac{-283.0 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{28.01 \text{ g}}$$

$\Delta H_c{}^\circ{}_{\text{per gram } CO(g)} = -10.10$ kJ/g

**Statement:** The enthalpy change per gram for the combustion of methane is –55.50 kJ/g and that of carbon monoxide is –10.10 kJ/g.

**(c)** Answers may vary. Sample answer: Methane is a better energy source as it releases 5 times more energy per gram than carbon monoxide.

**(d)** Answers may vary.

**Reflect on Your Learning**

**69.** Answers may vary. Sample answer: I was surprised to know these three pieces of information that I learned from this chapter:
• The average bond energies can be used for predicting enthalpy changes in chemical reactions.
• The standard enthalpies of formation can be used to calculate enthalpy change of a reaction.
• When a reaction proceeds from reactants to products, the change in enthalpy is the same whether the reaction occurs in one step or several.
I thought all enthalpy changes of reactions were not related, and they were supposed to be determined by measurements or experiments.

**70.** Answers may vary. Students' answers should include examples in everyday life of estimating the amount of energy released by substances, such a cold pack, ice, or alcohol, using standard enthalpies of formation or bond energies.

**71.** Answers may vary. Students' answers may include concepts such as solving problems using Hess's law, writing formation reactions, or writing combustion reactions and calculating enthalpy changes using standard enthalpy of formation values. They should improve their understanding of these concepts by making study notes, using graphic organizers to summarize information, or reviewing solutions to tutorial questions.

**72.** Answers may vary. Students' answers to misconceptions may include:

• Reactions represented by chemical equations cannot be added.

• Enthalpy changes of reactions are not related and are supposed to be determined by measurements or experiments.

**Research**

**73.** Answers may vary. Sample Answer:

|  | **Wind Farms** | **Solar Farms** |
|---|---|---|
| Advantages | • After construction, do not produce green house gases or pollutants<br>• Do not take up much space on land<br>• Energy can be captured efficiently with new technology<br>• Available in a variety of sizes<br>• Work for people who are not connected to a power grid | • Available in a variety of sizes<br>• Can be built near urban areas with little disruption<br>• Has maximum outputs on a hot day<br>• Power output can be adjusted quickly |
| Disadvantages | • Strength of wind not constant, so energy produced will vary<br>• Cause noise pollution<br>• Manufacturing wind turbines may cause some pollution, so wind power is not 100 % clean<br>• Provides less energy than common power stations; need many wind turbines<br>• People want to leave countryside intact, some consider them an eyesore | • Restricted by location, weather, time of day and year<br>• System has to be customized for different locations<br>• Need large surface area<br>• Manufacturing photovoltaic cells will produce harmful by-products and require large amounts of silicon<br>• Solar PV not as efficient as it could be; only about 12 to 17 % |
| Costs | • Raw fuel (wind) is free<br>• Setting up and maintenance are main costs<br>• Cost about 13.5 cents per KWh for on shore and 19 cents per KWh for off shore producers | • Raw fuel (sunshine) is free<br>• Solar PV farms currently cost around 15 cents per KWh<br>• High setting up and maintenance costs |

I do not agree with the environmental activist. I think that a combination of wind farms and solar farms would be best because some of the disadvantages of either one can be balanced out by the advantages of the other. This is especially true in developing countries where wind and solar energy can provide a continual, dependable source of electricity.

**74.** Answers may vary. Sample answers:

Natural Gas power plant:

**(a)** There is a number of ways to convert natural gas into electricity with varying amounts of fuel efficiency. The steam generation units use natural gas to heat up water so the steam produced can turn a turbine. It has low fuel efficiency with only about 33 to 35 % of the thermal energy can be converted into electrical energy. The centralized gas turbines are even less efficient than the steam generation unit. They use hot gases produced from burning natural gas to turn the turbine. Natural Gas-fired reciprocating engines or combustion turbines are more efficient with a percentage that ranges from 35 to 45 %. It turns the energy in natural gas into mechanical energy, which is then used to rotate a piston. Combined cycle units are even better with thermal efficiencies of up to 50 to 60 %. It uses both a gas turbine and a steam unit. The waste from using a gas turbine is used to produce steam for the steam unit. The most efficient type is microturbines. These are small industrial gas turbines that are about the size of a refrigerator. These can be up to 80 % efficient.

**(b)** Most of these methods involve burning natural gas. Natural gas is mainly composed of methane. The chemical reaction that produces the energy is the combustion of methane:

$2 \ CH_4(g) + 3 \ O_2(g) \rightarrow 2 \ CO_2(g) + 4 \ H_2O(l)$
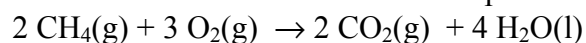
**(c)** The energy change for this energy source is –890 kJ/mol, or –55.5 kJ/g.

**75.** Answers may vary. Sample answers:

**(a)** The relationship between joules and calories is that while both measure energy, both are different in the way they measure energy. A calorie is defined as a unit of heat (which is a form of energy). It is the amount of energy needed to raise the temperature of 1 g of water by 1 °C. A joule is defined, as the amount of energy required to apply a force of 1 N over a distance of 1 m. One calorie is equal to 4.184 J. In addition, the calorie used to measure the amount of energy in food is different from the thermodynamic unit of a calorie. The one commonly used to represent energy in food is actually a kilocalorie, denoted by Cal while a thermodynamic calorie is denoted by cal. 1 Cal or kcal is $10^3$ cal.

The amount of calories in a 100 g plain bagel is 255 kcal.

The amount of calories in a 100 g piece of lean and fat beef is about 345 kcal.

The meat has a higher energy content then that of the bagel.

**(b)** The heat of combustion for octane is –48.23 kJ/g.

The heat of combustion for butane is –49.59 kJ/g.

Both reactions are exothermic.

Energy released with 100 g.

Octane: 48.23 kJ/g × 100 g = 48 230 kJ

Butane: 49.59 kJ/g × 100 g = 49 590 kJ

Bagel: 255 kcal × 4.184 kJ/kcal = 1 067 kJ

Beef: 345 kcal × 4.184 kJ/kcal = 1 443 kJ

Overall, it is clear that the hydrocarbons release more energy than food.

**76.** Answers may vary. Sample answers:

**(a)** The average amount of energy typically used in a Canadian home in 2007 is approximately 106 GJ or $106 \times 10^9$ J. The average amount of electrical energy is about 38% of this total, or 40.28 GJ.

**(b)** This quantity will vary depending on a number of factors such as size of solar panel, efficiency, insolation, and installed capacity.

For example:

Installed capacity: 15 W

Efficiency: 13 %

Insolation: 2 h

Convert GJ to kWh:

$40.28 \text{ GJ} = 4.028 \times 10^{10} \text{ J}$

$$4.028 \times 10^{10} \text{ J} \times \frac{1 \text{ kWh}}{3.6 \times 10^{6} \text{ J}} = 1.119 \times 10^{4} \text{ kWh}$$

Number of kWh produced per day:

$$\frac{1.119 \times 10^{4} \text{ kWh}}{365 \text{ days}} = 30.65 \text{ kWh}$$

Divide by insolation value:

$$\frac{30.65 \text{ kWh}}{2 \text{ h}} = 15.33 \text{ kW}$$

Factor in 13 % efficiency (energy losses and inefficiencies)

$15.33 \text{ kW} \times 1.13 = 17.32 \text{ kW}$

Therefore, a 17.32 kW system is needed to meet the energy demand.

Number of solar panels needed:

$$17.32 \text{ kW} \times \frac{1000 \text{ W}}{1 \text{ kW}} = 17 \; 320 \text{ W}$$

$$\frac{17 \; 320 \text{ W}}{15 \text{ W per panel}} = 1150 \text{ panels}$$

If all the solar panels were operating at maximum capacity, the minimum amount of panels needed are 1150 panels.

**(c)** No, it is not a feasible option because the minimum amount of panels needed is more than can be installed on a house.

**77.** Answers may vary. Sample answers:

**(a)** Hess did calorimetric experiments to study the different enthalpies exhibited by different chemical reactions. He used an ice calorimeter to conduct his experiments. By studying the changes in heat for various sulfuric acid hydrates as well as a variety of neutralization reaction, he concluded that the amount of heat was the same, no matter the pathway the reaction took.

**(b)** One limitation he had to deal with was the accuracy of the equipment. If his equipment were more accurate, his results would have been more accurate. One other limitation at that time was the information that was available. Hess was uncertain if the heat he measured was from the vibrational movement of particles or due to the physical properties of that substance.

**(c)** If Hess were working today the equipment he would use would be different. Nowadays, calorimeters are much more efficient and better at acting as isolated systems. In addition, an electrical thermometer can be used to measure the changes in heat. Electrical thermometers are more accurate than standard thermometers because they do not rely on the human eye to interpret the display.

**78.** Answers may vary. Sample answers:

**(a)** The history of the hybrid car extends all the way back in 1901. The first hybrid car, called the Mixte, was designed and developed by Ferdinand Porsche. It relied on a gasoline engine with an electrical engine giving it a boost of power. This method is still used to enhance performance and fuel economy. In 1915, the Woods Motor Vehicle manufacturers created the Dual Power hybrid vehicle. Instead of consolidating the two power sources to give a single source of power, it was designed so that an electric battery motor was used for low speeds and a gasoline engine for higher speeds. The Regenerative Braking System that was first developed by David Arthurs came in 1978. The electric motor is used to slow down the car and that power was used to help recharge the car batteries. Before this was conceived, a large amount of energy was lot as heat when the brakes were used in a conventional car. In the 1990s, Toyota and Honda became prominent in developing hybrid vehicles with the two first mainstream hybrid vehicle lines called the Toyota Prius and the Honda Insight.

Other forms of hybrid transportation include hybrid buses, hybrid trains and hybrid boats.

**(b)**

| Advantages | Disadvantages |
|---|---|
| • Lower emission and better fuel economy | • Cost more to buy compared to non-hybrid version |
| • Performs as well as gas-powered automobiles | • Weighs more due to batteries |
| • As reliable and comfortable as a conventional car | • Risk of high voltage wires during an accident |
| • Reduce need for fossil fuels and is better for environment | • Can only be fixed by experienced mechanic due to its complicated system |
| • Spend less on fuel | • Parts are expensive and hard to find |
| • Batteries do not need to be charged by external source | • Lower acceleration compared to non-hybrid version |
| • Special warranties alleviate cost of expensive parts | |

**(c)** I think a hybrid car is worth owning because the advantages outweigh the disadvantages. Although it costs more to buy initially, the rising cost of gas will make it more economic in the long run. In addition, the heavier weight is actually better for Canadian weather, especially during winter. Lastly, it will have less of a negative impact on the environment than buying a conventional vehicle.

**79.** Answers may vary. Sample answers:

**(a)** In 2010, the energy consumption is 7485.54 kg of oil equivalent per capita. China's energy consumption in 2009 is 1695.31 kg, which is lower than Canada's even though China has a much larger population. One trend in the data is that both countries are increasing their energy consumption with China at a very fast rate.

**(b)** Some factors are that both Canada and China's populations are increasing and the standard of living is increasing in both countries. In addition, China is a quickly developing country that is becoming more and more industrialized.

**(c)** Canada is not morally justified to continue this present level of energy consumption. We take more energy and generate more pollution to produce the same amount of goods and services in comparison to other competitors. In addition, Canada has one of the worst overall environmental records based on 25 environmental indicators. With countries like China who have a much higher population but use much less energy, and other developing nations in which basic amenities are not even met, are we justified in our careless consumption of resources?

**80.** Answers may vary. Sample answer: The chemical term for baking soda is sodium bicarbonate and it is effective for putting out grease fires because it smothers them. A fire needs three things; fuel, oxygen, and a high temperature. Sodium bicarbonate inhibits a fire's access to oxygen by producing carbon dioxide gas. Sodium bicarbonate decomposes into carbon dioxide, sodium carbonate, and water when heated over
50 °C.

The chemical reaction is $2\ NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$

It is an endothermic reaction. This means that as it decomposes, it also absorbs some of the heat. Thus, it is effective for putting out grease fires by not only absorbing some of the heat from the fire but also by smothering the fire by producing carbon dioxide.