
6. THERMOCHEMISTRY

■ Solutions to Exercises

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

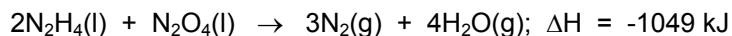
6.1 Substitute into the formula $E_k = 1/2 mv^2$ using SI units:

$$E_k = 1/2 \times 9.11 \times 10^{-31} \text{ kg} \times (5.0 \times 10^6 \text{ m/s})^2 = 1.\underline{13} \times 10^{-17} = 1.1 \times 10^{-17} \text{ J}$$

$$1.13 \times 10^{-17} \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 2.\underline{72} \times 10^{-18} = 2.7 \times 10^{-18} \text{ cal}$$

6.2 Heat is evolved; therefore, the reaction is exothermic. The value of q is -1170 kJ.

6.3 The thermochemical equation is

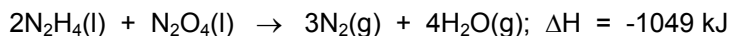


ΔH equals -1049 kJ for 1 mol N_2H_4 and for 2 mol N_2H_4 .

6.4 a. $\text{N}_2\text{H}_4(\text{l}) + 1/2\text{N}_2\text{O}_4(\text{l}) \rightarrow 3/2\text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}); \Delta H = -524.5 \text{ kJ}$

b. $4\text{H}_2\text{O}(\text{g}) + 3\text{N}_2(\text{g}) \rightarrow 2\text{N}_2\text{H}_4(\text{l}) + \text{N}_2\text{O}_4(\text{l}); \Delta H = 1049 \text{ kJ}$

6.5 The reaction is



$$10.0 \text{ g N}_2\text{H}_4 \times \frac{1 \text{ mol N}_2\text{H}_4}{32.02 \text{ g}} \times \frac{1 \text{ mol N}_2\text{O}_4}{2 \text{ mol N}_2\text{H}_4} \times \frac{-1049 \text{ kJ}}{1 \text{ mol N}_2\text{H}_4} = -163.80 = -164 \text{ kJ}$$

6.6 Substitute into the equation $q = s \times m \times \Delta t$ to obtain the heat transferred. The temperature change is

$$\Delta t = t_f - t_i = 100.0^\circ\text{C} - 20.0^\circ\text{C} = 80.0^\circ\text{C}$$

Therefore,

$$q = s \times m \times \Delta t = 0.449 \text{ J/(g}\cdot^\circ\text{C)} \times 5.00 \text{ g} \times 80.0^\circ\text{C} = 1.796 \times 10^2 = 1.80 \times 10^2 \text{ J}$$

6.7 The total mass of the solution is obtained by adding the volumes together and by using the density of water (1.000 g/mL). This gives $33 + 42 = 75 \text{ mL}$, or 75 g . The heat absorbed by the solution is

$$q = s \times m \times \Delta t = 4.184 \text{ J/(g}\cdot^\circ\text{C)} \times 75 \text{ g} \times (31.8^\circ\text{C} - 25.0^\circ\text{C}) = 2133.8 \text{ J}$$

The heat released by the reaction, q_{rxn} , is equal to the negative of this value, or -2133.8 J . To obtain the enthalpy change for the reaction, you need to calculate the moles of HCl that reacted. This is

$$\text{Mol HCl} = 1.20 \text{ mol/L} \times 0.033 \text{ L} = 0.0396 \text{ mol}$$

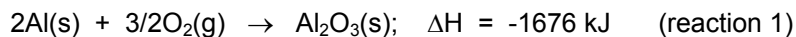
The enthalpy change for the reaction can now be calculated.

$$\Delta H = \frac{-2133.8 \text{ J}}{0.0396 \text{ mol}} = -53884 \text{ J/mol} = -54 \text{ kJ/mol}$$

Expressing this result as a thermochemical equation,



- 6.8 Use Hess's law to find ΔH for $4\text{Al(s)} + 3\text{MnO}_2\text{(s)} \rightarrow 2\text{Al}_2\text{O}_3 + 3\text{Mn(s)}$ from the following data for reactions 1 and 2:



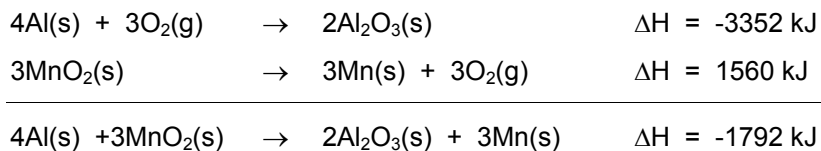
If you take reaction 1 and multiply it by 2, you obtain



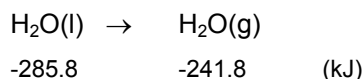
Since the desired reaction has three MnO_2 on the left side, reverse reaction 2 and multiply it by three. The result is



If you add the two reactions and corresponding enthalpy changes, you obtain the enthalpy change of the desired reaction



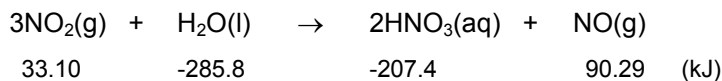
- 6.9 The vaporization process, with the ΔH_f° values given below the substances, is



The calculation is

$$\begin{aligned}
 \Delta H_{\text{vap}}^\circ &= \sum n \Delta H_f^\circ (\text{products}) - \sum m \Delta H_f^\circ (\text{reactants}) \\
 &= \Delta H_f^\circ [\text{H}_2\text{O(g)}] - \Delta H_f^\circ [\text{H}_2\text{O(l)}] \\
 &= (-241.8 \text{ kJ}) - (-285.8 \text{ kJ}) = 44.0 \text{ kJ}
 \end{aligned}$$

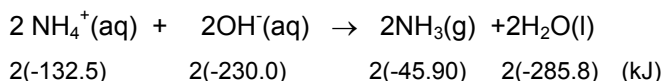
6.10 The reaction, with the ΔH_f° values given below the substances, is



The calculation is

$$\begin{aligned}
 \Delta H_{\text{rxn}}^\circ &= \sum n \Delta H_f^\circ (\text{products}) - \sum m \Delta H_f^\circ (\text{reactants}) \\
 &= [2 \Delta H_f^\circ (\text{HNO}_3) + \Delta H_f^\circ (\text{NO})] - [3 \Delta H_f^\circ (\text{NO}_2) + \Delta H_f^\circ (\text{H}_2\text{O})] \\
 &= [2(-207.4) + (90.29)] \text{ kJ} - [3(33.10) + (-285.8)] \text{ kJ} = -138.\underline{01} = -138.0 \text{ kJ}
 \end{aligned}$$

6.11 The net chemical reaction, with the ΔH_f° values given below the substances, is



The calculation is

$$\begin{aligned}
 \Delta H_{\text{rxn}}^\circ &= [2 \Delta H_f^\circ (\text{NH}_3) + 2 \Delta H_f^\circ (\text{H}_2\text{O})] - [2 \Delta H_f^\circ (\text{NH}_4^+) + 2 \Delta H_f^\circ (\text{OH}^-)] \\
 &= [2(-45.90) + 2(-285.8)] - [2(-132.5) + 2(-230.0)] = 61.\underline{60} = 61.6 \text{ kJ}
 \end{aligned}$$

■ Answers to Concept Checks

- 6.1 The photovoltaic cells collect the sun's energy, converting it to electric energy. This electric energy is stored in the battery as chemical energy, which is later changed back to electric energy that runs a motor. As the motor rotates, it changes the electric energy to kinetic energy (energy of motion) of the motor, then of water, which in turn is changed to potential energy (energy of position) of water as the water moves upward in the gravitational field of earth.
- 6.2 a. This reaction is the one shown in the problem, and it has a positive ΔH , so the reaction is endothermic.
- b. This reaction is simply twice reaction a, so it is also endothermic.
- c. This reaction is the reverse of reaction a, so it is exothermic.

(continued)

d. This reaction is simply twice that of reaction c, so it is more exothermic than c.

Thus, d is the most exothermic reaction.

- 6.3 You can think of the sublimation of ice as taking place in two stages. First, the solid melts to liquid, then the liquid vaporizes. The first process has an enthalpy, ΔH_{fus} . The second process has an enthalpy, ΔH_{vap} . Therefore, the total enthalpy, which is the enthalpy of sublimation, is the sum of these two enthalpies:

$$\Delta H_{\text{sub}} = \Delta H_{\text{fus}} + \Delta H_{\text{vap}}.$$

■ Answers to Review Questions

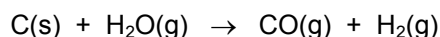
- 6.1 Energy is the potential or capacity to move matter. Kinetic energy is the energy associated with an object by virtue of its motion. Potential energy is the energy an object has by virtue of its position in a field of force. Internal energy is the sum of the kinetic and potential energies of the particles making up a substance.
- 6.2 In terms of SI base units, a joule is $\text{kg} \cdot \text{m}^2/\text{s}^2$. This is equivalent to the units for the kinetic energy ($\frac{1}{2}mv^2$) of an object of mass m (in kg) moving with speed v (in m/s).
- 6.3 Originally, a calorie was defined as the amount of energy required to raise the temperature of one gram of water by one degree Celsius. At present, the calorie is defined as 4.184 J.
- 6.4 At either of the two highest points above the earth in a pendulum's cycle, the energy of the pendulum is all potential energy and is equal to the product mgh (m = mass of pendulum, g = constant acceleration of gravity, and h = height of pendulum). As the pendulum moves downward, its potential energy decreases from mgh to near zero, depending on how close it comes to the earth's surface. During the downward motion, its potential energy is converted to kinetic energy. When it reaches the lowest point (middle) of its cycle, the pendulum has its maximum kinetic energy and minimum potential energy. As it rises above the lowest point, its kinetic energy begins to be converted to potential energy. When it reaches the other high point in its cycle, the energy of the pendulum is again all potential energy. By the law of conservation of energy, this energy cannot be lost, only converted to other forms. At rest, the energy of the pendulum has been transferred to the surroundings in the form of heat.
- 6.5 As the heat flows into the gas, the gas molecules gain energy and move at a faster average speed. The internal energy of the gas increases.

- 6.6 An exothermic reaction is a chemical reaction or a physical change in which heat is evolved (q is negative). For example, burning one mol of methane, $\text{CH}_4(\text{g})$, yields carbon dioxide, water, and 890.3 kJ of heat. An endothermic reaction is a chemical reaction or physical change in which heat is absorbed (q is positive). For example, the reaction of one mol of barium hydroxide with ammonium nitrate absorbs 170.4 kJ of heat in order to form ammonia, water, and barium nitrate.
- 6.7 Changes in internal energy depend only on the initial and final states of the system, which are determined by variables such as temperature and pressure. Such changes do not depend on any previous history of the system.
- 6.8 The enthalpy change equals the heat of reaction at constant pressure.
- 6.9 At constant pressure, the enthalpy change is positive (the enthalpy increases) for an endothermic reaction.
- 6.10 It is important to give the states when writing an equation for ΔH because ΔH depends on the states of all reactants and products. If any state changes, ΔH changes.
- 6.11 When the equation for the reaction is doubled, the enthalpy is also doubled. When the equation is reversed, the sign of ΔH is also reversed.
- 6.12 First, convert the 10.0 g of water to moles of water, using its molar mass (18.02 g/mol). Next, using the equation, multiply the moles of water by the appropriate mole ratio (1 mol CH_4 / 2 mol H_2O). Finally, multiply the moles of CH_4 by the heat of the reaction (-890.3 kJ/mol CH_4).
- 6.13 The heat capacity (C) of a substance is the quantity of heat needed to raise the temperature of the sample of substance one degree Celsius (or one kelvin). The specific heat of a substance is the quantity of heat required to raise the temperature of one gram of a substance by one degree Celsius (or one kelvin) at constant pressure.
- 6.14 A simple calorimeter consists of an insulated container (for example, a pair of styrene coffee cups as in Figure 6.12) with a thermometer. The heat of the reaction is obtained by conducting the reaction in the calorimeter. The temperature of the mixture is measured before and after the reaction. The heat capacity of the calorimeter and its contents also must be measured.
- 6.15 Hess's law states that, for a chemical equation that can be written as the sum of two or more steps, the enthalpy change for the overall equation is the sum of the enthalpy changes for the individual steps. In other words, no matter how you go from reactants to products, the enthalpy change for the overall chemical change is the same. This is because enthalpy is a state function.
- 6.16 No, you can still obtain the enthalpy for the desired reaction. You will need to come up with a system of reactions that can be combined to give the desired reaction and know the enthalpy changes for each of the steps. Then, using Hess's law, you can obtain the enthalpy change for the reaction under study.

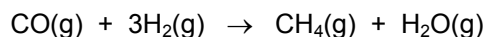
- 6.17 The thermodynamic standard state refers to the standard thermodynamic conditions chosen for substances when listing or comparing thermodynamic data: one atm pressure and the specified temperature (usually 25°C).
- 6.18 The reference form of an element is the most stable form (physical state and allotrope) of the element under standard thermodynamic conditions. The standard enthalpy of formation of an element in its reference form is zero.
- 6.19 The standard enthalpy of formation of a substance, ΔH°_f , is the enthalpy change for the formation of one mole of the substance in its standard state from its elements in their reference form and in their standard states.
- 6.20 The equation for the formation of $\text{H}_2\text{S}(\text{g})$ is



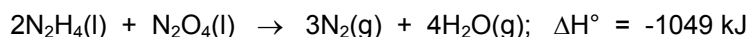
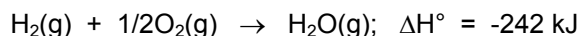
- 6.21 The reaction of $\text{C}(\text{g}) + 4\text{H}(\text{g}) \rightarrow \text{CH}_4(\text{g})$ is not an appropriate equation for calculating the ΔH°_f of methane because the most stable form of each element is not used. Both $\text{H}_2(\text{g})$ and $\text{C}(\text{graphite})$ should be used instead of $\text{H}(\text{g})$ and $\text{C}(\text{g})$, respectively.
- 6.22 A fuel is any substance that is burned or similarly reacted to provide heat and other forms of energy. The fossil fuels are petroleum (oil), gas, and coal. They were formed millions of years ago when aquatic plants and animals were buried and compressed by layers of sediment at the bottoms of swamps and seas. Over time this organic matter was converted by bacterial decay and pressure to fossil fuels.
- 6.23 One of the ways of converting coal to methane involves the water-gas reaction.



In this reaction, steam is passed over hot coal. This mixture is then reacted over a catalyst to give methane.

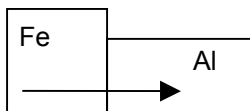


- 6.24 Some possible rocket fuel/oxidizer combinations are H_2/O_2 and hydrazine/dinitrogen tetroxide. The chemical equations for their reactions are



■ Answers to Conceptual Problems

- 6.25 Kinetic energy is proportional to mass and to speed squared. Compare the kinetic energy of the smaller car with that of the larger car (twice the mass) assuming both are traveling at the same speed. The larger car would have twice the kinetic energy of the smaller car. Or, we could say that the smaller car has only half the kinetic energy of the larger car. Now suppose the speed of the smaller is increased by a factor of two (so it is now moving at twice its original speed). Its kinetic energy is increased by a factor of four. Therefore, the smaller car now has one-half times four times, or twice, the kinetic energy of the larger car. The smaller car has the greater kinetic energy.
- 6.26 The equation says that one mol of butane reacts with $(13/2)$ mol of oxygen to yield four mol carbon dioxide and five mol water. The reaction yields a certain amount of heat, which you can symbolize as q . So a yields heat q . On the other hand, b is only one mol oxygen, not $(13/2)$ mol. So, b yields heat equal to $(2/13) q$. This result might be easier to see by first looking at c . Note that the one mole of carbon dioxide stated in c is only one-fourth that given in the equation. This means that c yields one-fourth q (just the inverse of the coefficient of the equation). Similarly, d yields one-fifth q . Therefore, b yields the least heat.
- 6.27 a. After the water is placed into the freezer, it will lose heat to the freezer, so q_{sys} is negative.
- b. The water will have turned to ice.
- c. The initial enthalpy (of the water) is higher than the final enthalpy (of the ice).
- d. After several hours, the temperature of the water will be -20°C .
- 6.28 a. Heat will flow from the iron block to the aluminum block.



- b. The aluminum block will be absorbing heat, so q_{sys} will be positive.
- c. The temperatures of the aluminum block and the iron block will both be the same.
- d. The heat gained by the aluminum will equal the heat lost by the iron. Each heat term will be equal to mass \times sp ht $\times \Delta T$. This gives

$$(20.0 \text{ g})(0.901 \text{ J/g}^{\circ}\text{C})(T_f - 45.0^{\circ}\text{C}) = (20.0 \text{ g})(0.449 \text{ J/g}^{\circ}\text{C})(50.0^{\circ}\text{C} - T_f)$$

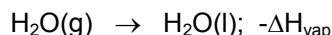
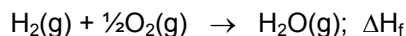
(continued)

You can cancel the mass and the units from both sides to give a simplified expression

$$(0.901)(T_f - 45.0^\circ\text{C}) = (0.449)(50.0^\circ\text{C} - T_f)$$

This can be simplified to give $1.350 T_f = 62.995$, or $T_f = 46.\underline{66} = 46.7^\circ\text{C}$.

- 6.29 You can imagine this process taking place in two steps: first, the preparation of water vapor from the elements, and second, the change of the vapor to liquid. Here are the equations:



The last equation is the reverse of the vaporization of water, so the enthalpy of the step is the negative of the enthalpy of vaporization. The enthalpy change for the preparation of one mole of liquid water, ΔH , is the sum of the enthalpy changes for these two steps:

$$\Delta H = \Delta H_f + (-\Delta H_{\text{vap}}) = \Delta H_f - \Delta H_{\text{vap}}$$

- 6.30 The expression for the heat is $q = s \times m \times \Delta t$. For the same amount of heat and mass, the product $s \times \Delta t$ must be constant. The metal with the smaller specific heat will have the larger Δt . Since the specific heat for aluminum ($0.901 \text{ J/g}^\circ\text{C}$) is larger than for iron ($0.449 \text{ J/g}^\circ\text{C}$), the block of iron will have the larger Δt and be warmer.
- 6.31 a. The heat lost by the metal is equal to the heat gained by the water. Since $q = s \times m \times \Delta t$, the heat gained by the water is directly proportional to Δt . Since Δt is larger for metal A, it lost more heat. Now, each metal has the same mass and Δt , so the specific heat is directly proportional to q . Since q is larger for A, the specific heat is larger for A.
- b. The metal with the higher specific heat will have absorbed more heat to reach the starting temperature of 95°C ; therefore, it will release more heat to the water causing the water to reach a higher temperature. The beaker with metal A will rise to the higher temperature.
- 6.32 Silver metal reacts with a halogen to produce the corresponding silver halide. For example, silver reacts with fluorine to produce silver fluoride. Each reaction corresponds to the reaction forming the silver halide, so you look in the table of enthalpies of formation of compounds (Table 6.2). The most exothermic reaction would be the one with the most negative enthalpy of formation. That would be the reaction for the formation of silver fluoride.

6.33 Let us write ΔH_{rxn} for the enthalpy change when one mole of P_4S_3 burns in O_2 to give P_4O_{10} and SO_2 . In principle, you could calculate ΔH_{rxn} from enthalpies of formation for the reactants and products of this reaction. You would require the values for P_4S_3 , O_2 (which equals zero), P_4O_{10} , and SO_2 . Enthalpies of formation for the products, P_4O_{10} and SO_2 , are given in Table 6.2 in the text. This means that, if you have measured ΔH_{rxn} , you can use the enthalpies of formation of P_4O_{10} and SO_2 to calculate the enthalpy of formation for P_4S_3 . What you have done is this: you have used enthalpies of combustion to calculate enthalpies of formation. This is the idea most often used to obtain enthalpies of formation.

- 6.34 a. Since ΔH is positive, the reaction is endothermic, and the solution will be colder.
- b. While the salt dissolves, heat will flow into the beaker to raise the temperature of the water back to the initial temperature.
- c. After the water returns to room temperature, q for the system will be zero.

■ Solutions to Practice Problems

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

6.35 The heat released, in kcal, is

$$-445.1 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = -106.\underline{38} = -106.4 \text{ kcal}$$

6.36 The heat released, in kcal, is

$$-281.0 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = -67.1\underline{60} = -67.16 \text{ kcal}$$

6.37 The kinetic energy, in J, is

$$E_k = 1/2 \times 4.85 \times 10^3 \text{ lb} \times \frac{0.4536 \text{ kg}}{1 \text{ lb}} \times \left[\frac{57 \text{ mi}}{1 \text{ h}} \right]^2 \times \left[\frac{1609 \text{ m}}{1 \text{ mi}} \right]^2 \times \left[\frac{1 \text{ h}}{3600 \text{ s}} \right]^2$$

$$= 7.13 \times 10^5 = 7.1 \times 10^5 \text{ J}$$

The kinetic energy, in calories, is

$$E_k = 7.13 \times 10^5 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 1.70 \times 10^5 = 1.7 \times 10^5 \text{ cal}$$

6.38 The kinetic energy, in J, is

$$E_k = 1/2 \times 235 \text{ gr} \times \frac{0.0000648 \text{ kg}}{1 \text{ gr}} \times \left[\frac{2520 \text{ ft}}{1 \text{ s}} \right]^2 \times \left[\frac{1 \text{ yd}}{3 \text{ ft}} \right]^2 \times \left[\frac{0.9144 \text{ m}}{1 \text{ yd}} \right]^2$$

$$= 4492 = 4.49 \times 10^3 \text{ J}$$

The kinetic energy, in cal, is

$$E_k = 4492 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 1073 = 1.07 \times 10^3 \text{ cal}$$

6.39 To insert the mass of one molecule of ClO_2 in the formula, multiply the molar mass by the reciprocal of Avogadro's number. The kinetic energy in J is

$$E_k = 1/2 \times \frac{67.45 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molec.}} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \left[\frac{306 \text{ m}}{1 \text{ s}} \right]^2$$

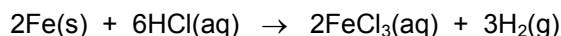
$$= 5.243 \times 10^{-21} = 5.24 \times 10^{-21} \text{ J/molec.}$$

6.40 To insert the mass of one molecule of N_2O in the formula, multiply the molar mass by the reciprocal of Avogadro's number. The kinetic energy in J is

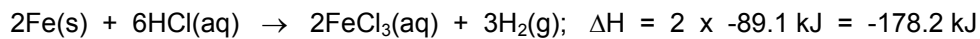
$$E_k = 1/2 \times \frac{44.02 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molec.}} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \left[\frac{379 \text{ m}}{1 \text{ s}} \right]^2$$

$$= 5.249 \times 10^{-21} = 5.25 \times 10^{-21} \text{ J/molec.}$$

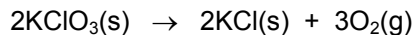
- 6.41 Endothermic reactions absorb heat, so the sign of q will be positive because energy must be gained by the system from the surroundings. The flask will feel cold to the touch.
- 6.42 Exothermic reactions evolve heat, so the sign of q will be negative because energy is lost from the system to the surroundings. The flask will feel hot to the touch.
- 6.43 The gain of 66.2 kJ of heat per two mol NO_2 means the reaction is endothermic. Because energy is gained by the system from the surroundings, q is positive and is +66.2 kJ for two mol of NO_2 reacting.
- 6.44 The release of 939 kJ of heat per two mol HCN means the reaction is exothermic. Because energy is lost from the system to the surroundings, q is negative, and because the reaction involves two mol HCN , q for the reaction is -939 kJ.
- 6.45 The reaction of Fe(s) with HCl must yield H_2 and FeCl_3 . To balance the hydrogen, 2HCl must be written first as a reactant. However, to balance the chlorine, $3 \times 2\text{HCl}$ must be written finally as a reactant, and 2FeCl_3 must be written as a product:



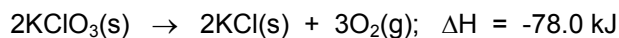
To write a thermochemical equation, the sign of ΔH must be negative because heat is evolved:



- 6.46 The decomposition of two mol KClO_3 to KCl and O_2 must yield two mol KCl to balance the potassium and chlorine. To balance the oxygen, three mol O_2 must be written as a product:



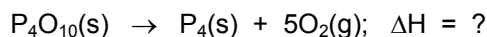
To write a thermochemical equation, the sign of ΔH must be negative because heat is evolved:



6.47 The first equation is

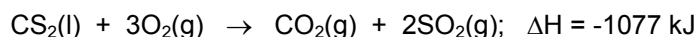


The second equation is

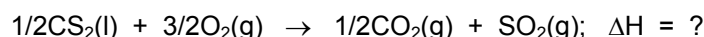


The second equation has been obtained by reversing the first equation. Therefore, to obtain ΔH for the second equation, ΔH for the first equation must be reversed in sign: $-(-3010) = +3010 \text{ kJ}$.

6.48 The first equation is

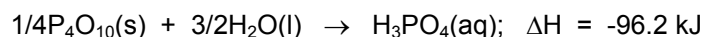


The second equation is

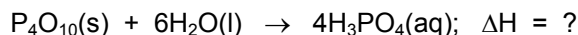


The second equation has been obtained by dividing each coefficient in the first equation by two. Therefore, to obtain ΔH for the second equation, ΔH for the first equation must be divided by two: $(-1077) \div 2 = -538.5 = -539 \text{ kJ}$.

6.49 The first equation is

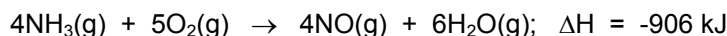


The second equation is

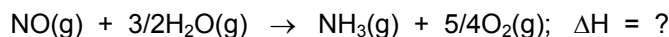


The second equation has been obtained from the first by multiplying each coefficient by four. Therefore, to obtain the ΔH for the second equation, the ΔH for the first equation must be multiplied by four: $-96.2 \times 4 = -384.8 \text{ kJ}$

6.50 The first equation is



The second equation is



The second equation has been obtained by reversing the first equation and dividing each coefficient by four. Therefore, to obtain ΔH for the second equation, the ΔH for the first equation must be reversed in sign and divided by four: $+906 \div 4 = 226.50 = 227 \text{ kJ}$.

- 6.51 Because nitric oxide is written as NO in the equation, the molar mass of NO equals 30.01 g per mol NO. From the equation, two mol NO evolve 114 kJ heat. Divide the 114 kJ by the 30.01 g/mol NO to obtain the amount of heat evolved per gram of NO:

$$\frac{-114 \text{ kJ}}{2 \text{ mol NO}} \times \frac{1 \text{ mol NO}}{30.01 \text{ g NO}} = 1.899 = \frac{-1.90 \text{ kJ}}{\text{g NO}}$$

- 6.52 Because hydrogen is written as H₂ in the equation, the molar mass of H₂ equals 2.016 g per mol H₂. From the equation, two mol H₂ evolve 484 kJ heat. Divide this by the 2.016 g/mol H₂ to obtain the amount of heat evolved per gram of hydrogen:

$$\frac{-484 \text{ kJ}}{2 \text{ mol H}_2} \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} = -120.03 = \frac{-120 \text{ kJ}}{\text{g H}_2}$$

- 6.53 The molar mass of ammonia is 17.03 g/mol. From the equation, four moles of NH₃ evolve 1267 kJ of heat. Divide 35.8 g NH₃ by its molar mass and the four moles of NH₃ in the equation to obtain the amount of heat evolved:

$$35.8 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{-1267 \text{ kJ}}{4 \text{ mol NH}_3} = -6.658 \times 10^2 = -6.66 \times 10^2 \text{ kJ}$$

- 6.54 The molar mass of H₂S is 34.08 g/mol. From the equation, two moles of H₂S evolve 1036 kJ of heat. This gives

$$26.7 \text{ g H}_2\text{S} \times \frac{1 \text{ mol H}_2\text{S}}{34.09 \text{ g H}_2\text{S}} \times \frac{-1036 \text{ kJ}}{2 \text{ mol H}_2\text{S}} = -405.7 = -406 \text{ kJ}$$

- 6.55 The molar mass of C₃H₈ is 44.06 g/mol. From the equation, one mol C₃H₈ evolves 2043 kJ heat. This gives

$$-369 \text{ kJ} \times \frac{1 \text{ mol C}_3\text{H}_8}{-2043 \text{ kJ}} \times \frac{44.06 \text{ g C}_3\text{H}_8}{1 \text{ mol C}_3\text{H}_8} = 7.957 = 7.96 \text{ g C}_3\text{H}_8$$

- 6.56 The molar mass of C₂H₅OH is 46.05 g/mol. From the equation, one mol C₂H₅OH evolves 1235 kJ heat. This gives

$$-293 \text{ kJ} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{-1235 \text{ kJ}} \times \frac{46.05 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 10.92 = 10.9 \text{ g C}_2\text{H}_5\text{OH}$$

- 6.57 Multiply the 180 g (0.180 kg) of water by the specific heat of 4.18 J/(g•°C) and by Δt to obtain heat in joules:

$$180 \text{ g} \times (96^\circ\text{C} - 19^\circ\text{C}) \times \frac{4.18 \text{ J}}{1 \text{ g}\cdot^\circ\text{C}} = 57934 = 5.8 \times 10^4 \text{ J}$$

- 6.58 Multiply the 1.63×10^3 g (1.63 kg) iron by the specific heat of 0.450 J/(g•°C) and by Δt to obtain heat in joules:

$$1.63 \times 10^3 \text{ g} \times (178^\circ\text{C} - 21^\circ\text{C}) \times \frac{0.449 \text{ J}}{1 \text{ g}\cdot^\circ\text{C}} = 114903 = 1.15 \times 10^5 \text{ J}$$

- 6.59 Use the 2.26×10^3 J/g (2.26 kJ/g) heat of vaporization to calculate the heat of condensation. Then, use it to calculate Δt , the temperature change.

$$\text{Heat of condensation} = \frac{2.26 \times 10^3 \text{ J}}{1 \text{ g}} \times 168 \text{ g} = 3.7968 \times 10^5 \text{ J}$$

$$\text{Temperature change} = \Delta t = \frac{3.7968 \times 10^5 \text{ J}}{6.44 \times 10^4 \text{ g}} \times \frac{1 \text{ g}\cdot^\circ\text{C}}{1.015 \text{ J}} = 5.8085 = 5.81^\circ\text{C}$$

- 6.60 Use the 334 J/g (0.334 kJ/g) heat of fusion to calculate the heat of melting. Then, use the heat of melting to calculate Δt , the temperature change.

$$\text{Heat of melting} = 38.0 \text{ g} \times \frac{334 \text{ J}}{1 \text{ g}} = 12692 \text{ J}$$

$$\text{Heat melting} + \text{heat warming} = \text{heat lost by } 210 \text{ g H}_2\text{O}$$

Let T = the final temperature, and substitute into the above equation:

$$\begin{aligned} 12692 \text{ J} + \left[\frac{4.18 \text{ J}}{1 \text{ g}\cdot^\circ\text{C}} \times 38.0 \text{ g} \times (T - 0^\circ\text{C}) \right] \\ = \left[\frac{4.18 \text{ J}}{1 \text{ g}\cdot^\circ\text{C}} \times 210 \text{ g} \times (21.0^\circ\text{C} - T) \right] \end{aligned}$$

Solve the above equation for T .

$$T = \frac{18433.8 \text{ J} - 12692 \text{ J}}{(158.84 + 877.8) \text{ J}/^\circ\text{C}} = 5.538 = 5.54^\circ\text{C}$$

- 6.61 The enthalpy change for the reaction is equal in magnitude and opposite in sign to the heat-energy change occurring from the cooling of the solution and calorimeter.

$$q_{\text{calorimeter}} = (1071 \text{ J/}^{\circ}\text{C})(21.56^{\circ}\text{C} - 25.00^{\circ}\text{C}) = -3684.2 \text{ J}$$

Thus, 15.3 g NaNO_3 is equivalent to -3684.2 J heat energy. The amount of heat absorbed by 1.000 mol NaNO_3 is calculated from +3684.2 J (opposite sign):

$$1.000 \text{ mol NaNO}_3 \times \frac{85.00 \text{ g NaNO}_3}{1 \text{ mol NaNO}_3} \times \frac{3684.2 \text{ J}}{15.3 \text{ g NaNO}_3} = 2.0468 \times 10^4 \text{ J}$$

Thus, the enthalpy change, ΔH , for the reaction is $2.05 \times 10^4 \text{ J}$, or 20.5 kJ, per mol NaNO_3 .

- 6.62 The enthalpy change for the reaction is equal in magnitude and opposite in sign to the heat energy produced from the warming of the solution and the calorimeter.

$$q_{\text{calorimeter}} = (1258 \text{ J/}^{\circ}\text{C})(38.7^{\circ}\text{C} - 25.0^{\circ}\text{C}) = +1.7234 \times 10^4 \text{ J}$$

Thus, 23.6 g CaCl_2 is equivalent to $1.7234 \times 10^4 \text{ J}$ heat energy. The amount of heat released by 1.000 mol CaCl_2 is calculated from $-1.7234 \times 10^4 \text{ J}$ (opposite sign):

$$1.000 \text{ mol CaCl}_2 \times \frac{110.98 \text{ g CaCl}_2}{1 \text{ mol CaCl}_2} \times \frac{-1.7234 \times 10^4 \text{ J}}{23.6 \text{ g CaCl}_2} = -8.104 \times 10^4 \text{ J}$$

Thus, the enthalpy change, ΔH , for the reaction is $-8.10 \times 10^4 \text{ J}$, or -81.0 kJ, per mol CaCl_2 .

- 6.63 The energy change for the reaction is equal in magnitude and opposite in sign to the heat energy produced from the warming of the solution and the calorimeter.

$$q_{\text{calorimeter}} = (9.63 \text{ kJ/}^{\circ}\text{C})(33.73 - 25.00^{\circ}\text{C}) = +84.069 \text{ kJ}$$

Thus, 2.84 g $\text{C}_2\text{H}_5\text{OH}$ is equivalent to 84.069 kJ heat energy. The amount of heat released by 1.000 mol $\text{C}_2\text{H}_5\text{OH}$ is calculated from -84.069 kJ (opposite sign):

$$1.00 \text{ mol C}_2\text{H}_5\text{OH} \times \frac{46.07 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \times \frac{-84.069 \text{ kJ}}{2.84 \text{ g C}_2\text{H}_5\text{OH}} = -1363.75 \text{ kJ}$$

Thus, the enthalpy change, ΔH , for the reaction is $-1.36 \times 10^3 \text{ kJ/mol}$ ethanol.

- 6.64 The energy change for the reaction is equal in magnitude and opposite in sign to the heat energy produced from the warming of the solution and the calorimeter.

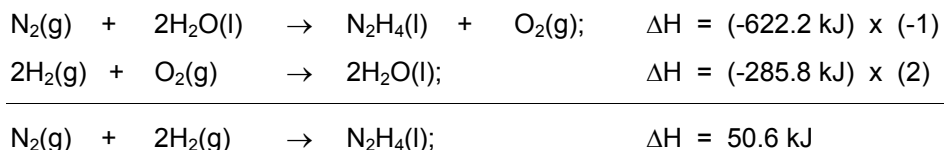
$$q_{\text{calorimeter}} = (12.05 \text{ kJ/}^\circ\text{C})(37.18^\circ\text{C} - 25.00^\circ\text{C}) = +146.769 \text{ kJ}$$

Thus, 3.51 g C_6H_6 is equivalent to 146.769 kJ heat energy. The amount of heat released by 1.000 mol C_6H_6 is calculated from -146.769 kJ (opposite sign):

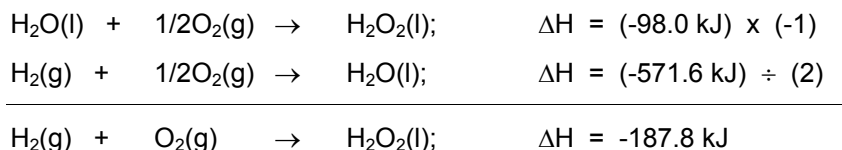
$$1.00 \text{ mol C}_6\text{H}_6 \times \frac{78.11 \text{ g C}_6\text{H}_6}{1 \text{ mol C}_6\text{H}_6} \times \frac{-146.769 \text{ kJ}}{3.51 \text{ g C}_6\text{H}_6} = -3266.1$$

Thus, the enthalpy change, ΔH , for the reaction is $-3.27 \times 10^3 \text{ kJ/mol}$ benzene.

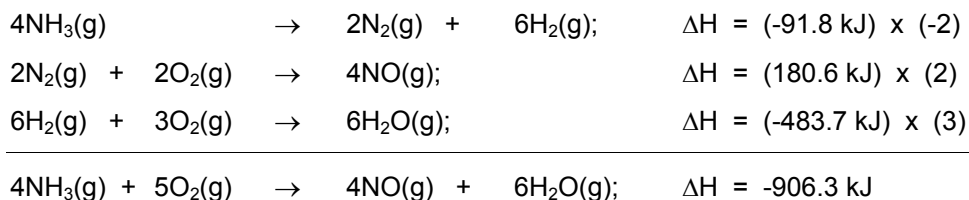
- 6.65 Using the equations in the data, reverse the direction of the first reaction, and reverse the sign of its ΔH . Then multiply the second equation by two, multiply its ΔH by two, and add. Setup:



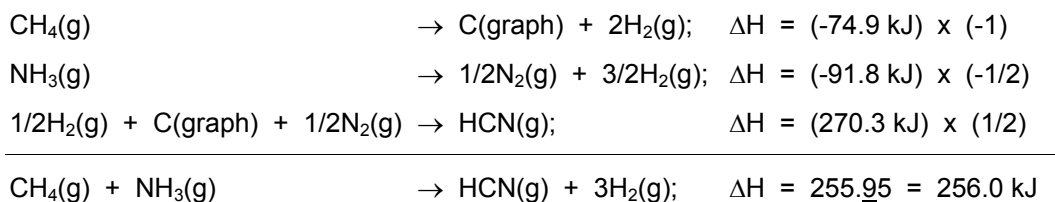
- 6.66 Using the equations in the data, reverse the direction of the first reaction, and reverse the sign of its ΔH . Then divide the second equation by two, divide its ΔH by two, and add. Setup:



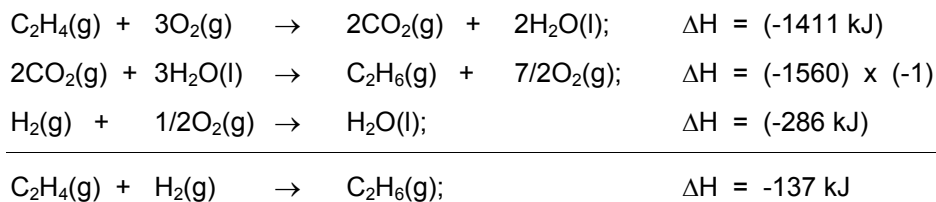
- 6.67 Using the equations in the data, multiply the second equation by two, and reverse its direction; do the same to its ΔH . Then multiply the first equation by two and its ΔH by two. Finally, multiply the third equation by three and its ΔH by three. Then add. Setup:



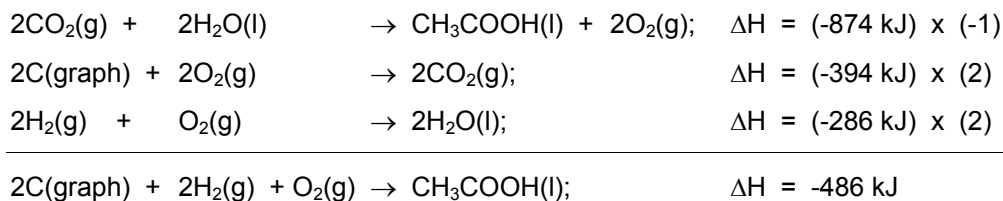
- 6.68 Using the equations in the data, reverse the direction of the second equation and reverse the sign of its ΔH . Then, reverse the direction of the first equation and multiply by $1/2$; do the same to its ΔH . Finally, multiply the third equation and its ΔH by $1/2$. Add all three equations. Setup:



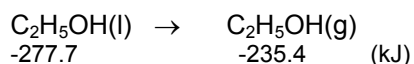
- 6.69 After reversing the first equation in the data, add all the equations. Setup:



- 6.70 After reversing the first equation, double the second and third equations, and then add. Setup:

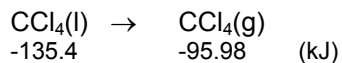


- 6.71 Write the ΔH° values (Appendix C) underneath each compound in the balanced equation:



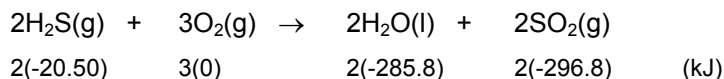
$$\Delta H^\circ_{\text{vap}} = [\Delta H^\circ_f(\text{C}_2\text{H}_5\text{OH})] - [\Delta H^\circ_f(\text{C}_2\text{H}_5\text{OH})] = [-235.4] - [-277.7] = +42.3 \text{ kJ}$$

- 6.72 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation:



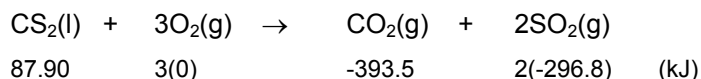
$$\Delta H^\circ_{\text{vap}} = [\Delta H^\circ_f(\text{CCl}_4)] - [\Delta H^\circ_f(\text{CCl}_4)] = [-95.98] - [-135.4] \text{ kJ} = +39.4 \text{ kJ}$$

- 6.73 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation:



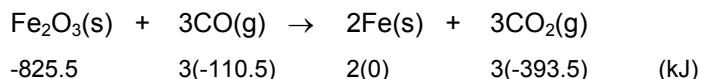
$$\begin{aligned} \Delta H^\circ &= \Sigma n\Delta H^\circ(\text{products}) - \Sigma m\Delta H^\circ(\text{reactants}) \\ &= [2(-285.8) + 2(-296.8)] - [2(-20.50) + 3(0)] \text{ kJ} = -1124.2 = -1124 \text{ kJ} \end{aligned}$$

- 6.74 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation:



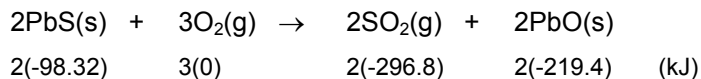
$$\begin{aligned} \Delta H^\circ &= \Sigma n\Delta H^\circ(\text{products}) - \Sigma m\Delta H^\circ(\text{reactants}) \\ &= [(-393.5) + 2(-296.8)] - [(87.90) + 3(0)] = -1075.0 \text{ kJ} \end{aligned}$$

- 6.75 Write the ΔH° values (Appendix C) underneath each compound in the balanced equation:



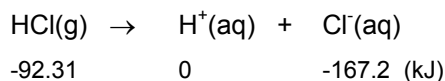
$$\begin{aligned} \Delta H^\circ &= \Sigma n\Delta H^\circ(\text{products}) - \Sigma m\Delta H^\circ(\text{reactants}) \\ &= [2(0) + 3(-393.5)] - [(-825.5) + 3(-110.5)] = -23.5 \text{ kJ} \end{aligned}$$

- 6.76 Write the ΔH° values (Appendix C) underneath each compound in the balanced equation:



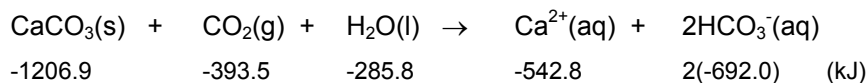
$$\begin{aligned} \Delta H^\circ &= \Sigma n\Delta H^\circ(\text{products}) - \Sigma m\Delta H^\circ(\text{reactants}) \\ &= [2(-296.8) + 2(-219.4)] - [2(-98.32) + 3(0)] = -835.76 = -835.8 \text{ kJ} \end{aligned}$$

- 6.77 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation:



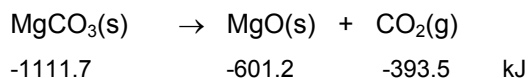
$$\begin{aligned} \Delta H^\circ &= \Sigma n\Delta H^\circ(\text{products}) - \Sigma m\Delta H^\circ(\text{reactant}) \\ &= [(0) + (-167.2)] - [-92.31] = -74.89 = -74.9 \text{ kJ} \end{aligned}$$

- 6.78 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation:



$$\begin{aligned} \Delta H^\circ &= \Sigma n\Delta H^\circ(\text{products}) - \Sigma m\Delta H^\circ(\text{reactants}) \\ &= [(-542.8) + 2(-692.0)] - [(-1206.9) + (-393.5) + (-285.8)] = -40.6 \text{ kJ} \end{aligned}$$

- 6.79 Calculate the molar heat of formation from the equation with the ΔH°_f values below each substance; then convert to the heat for 10.0 g of MgCO_3 using the molar mass of 84.3.

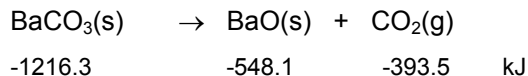


$$\Delta H^\circ = \Sigma n\Delta H^\circ(\text{products}) - \Sigma m\Delta H^\circ(\text{reactants})$$

$$\Delta H^\circ = -601.2 \text{ kJ} + (-393.5 \text{ kJ}) - (-1111.7 \text{ kJ}) = 117.0 \text{ kJ}$$

$$\text{Heat} = 10.0 \text{ g} \times \frac{1 \text{ mol}}{84.3 \text{ g}} \times \frac{117.0 \text{ kJ}}{\text{mol}} = 13.87 = 13.9 \text{ kJ}$$

- 6.80 Calculate the molar heat of formation from the equation with the ΔH°_f values below each substance; then convert to the heat for 10.0 g of BaCO_3 using the molar mass of 197.3.



$$\Delta H^\circ = \Sigma n\Delta H^\circ(\text{products}) - \Sigma m\Delta H^\circ(\text{reactants})$$

$$\Delta H^\circ = -548.1 \text{ kJ} + (-393.5 \text{ kJ}) - (-1216.3 \text{ kJ}) = 274.7 \text{ kJ}$$

$$\text{Heat} = 10.0 \text{ g} \times \frac{1 \text{ mol}}{197.3 \text{ g}} \times \frac{274.7 \text{ kJ}}{\text{mol}} = 13.92 = 13.9 \text{ kJ}$$

■ Solutions to General Problems

- 6.81 The SI units of force must be $\text{kg}\cdot\text{m}/\text{s}^2$ (= newton, N) to be consistent with the joule, the SI unit of energy:

$$\frac{\text{kg}\cdot\text{m}}{\text{s}^2} \times \text{m} = \frac{\text{kg}\cdot\text{m}^2}{\text{s}^2} = \text{joule, J}$$

- 6.82 Solving $E_p = mgh$, you obtain $g = E_p/mh$. Thus, the SI unit of g is $(\text{kg}\cdot\text{m}^2/\text{s}^2)/(\text{kg}\cdot\text{m}) = \text{m}/\text{s}^2$.

- 6.83 Using Table 1.5 and 4.184 J/cal, convert the 686 Btu/lb to J/g:

$$\begin{aligned} \frac{686 \text{ Btu}}{1 \text{ lb}} \times \frac{252 \text{ cal}}{1 \text{ Btu}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ lb}}{0.4536 \text{ kg}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \\ = 1.594 \times 10^3 = 1.59 \times 10^3 \text{ J/g} \end{aligned}$$

- 6.84 Using Table 1.5 and 4.184 J/cal, convert the 51600 Btu/lb to J/g:

$$\begin{aligned} \frac{51600 \text{ Btu}}{1 \text{ lb}} \times \frac{252 \text{ cal}}{1 \text{ Btu}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ lb}}{0.4536 \text{ kg}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \\ = 1.199 \times 10^5 = 1.20 \times 10^5 \text{ J/g} \end{aligned}$$

- 6.85 Substitute into the equation $E_p = mgh$, and convert to SI units.

$$\begin{aligned} E_p &= 1.00 \text{ lb} \times \frac{9.807 \text{ m}}{\text{s}^2} \times 167 \text{ ft} \times \frac{0.4536 \text{ kg}}{1 \text{ lb}} \times \frac{0.9144 \text{ m}}{3 \text{ ft}} \\ &= \frac{226.433 \text{ kg}\cdot\text{m}^2}{\text{s}^2} = 226 \text{ J} \end{aligned}$$

At the bottom, all the potential energy is converted to kinetic energy, so $E_k = 225.8 \text{ kg}\cdot\text{m}^2/\text{s}^2$. Because $E_k = 1/2mv^2$, solve for v , the speed (velocity):

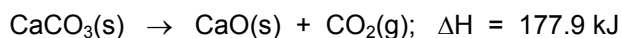
$$\text{Speed} = \sqrt{\frac{E_k}{1/2 \times m}} = \sqrt{\frac{226.433 \text{ kg}\cdot\text{m}^2/\text{s}^2}{1/2 \times 1.00 \text{ lb} \times 0.4536 \text{ kg/lb}}} = 31.60 = 31.6 \text{ m/s}$$

6.86 Substitute into the equation $E_k = 1/2mv^2$, and convert to SI units.

$$E_k = 1/2 \times 2354 \text{ lb} \times \frac{0.4536 \text{ kg}}{1 \text{ lb}} \times \left(\frac{11.2 \text{ km}}{\text{s}} \times \frac{10^3 \text{ m}}{1 \text{ km}} \right)^2$$

$$= 6.697 \times 10^{10} \text{ kg} \cdot \text{m}^2/\text{s}^2 = 6.70 \times 10^{10} \text{ J}$$

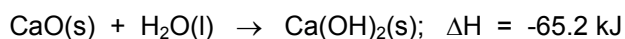
6.87 The equation is



Use the molar mass of 100.08 g/mol to convert the heat per mol to heat per 27.3 g.

$$21.3 \text{ g CaCO}_3 \times \frac{1 \text{ mol CaCO}_3}{100.08 \text{ g CaCO}_3} \times \frac{177.9 \text{ kJ}}{1 \text{ mol CaCO}_3} = 37.\underline{86} = 37.9 \text{ kJ}$$

6.88 The equation is

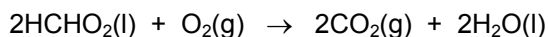


Use the molar mass in the conversion.

$$24.5 \text{ g CaO} \times \frac{1 \text{ mol CaO}}{56.08 \text{ g CaO}} \times \frac{-65.2 \text{ kJ}}{1 \text{ mol CaO}} = -28.\underline{48} = -28.5 \text{ kJ}$$

The heat released is 28.5 kJ.

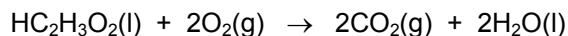
6.89 The equation is



Use the molar mass of 46.03 g/mol to convert -30.3 kJ/5.48 g to ΔH per mol of acid.

$$\frac{-30.3 \text{ kJ}}{5.48 \text{ g HCHO}_2} \times \frac{46.03 \text{ g HCHO}_2}{1 \text{ mol HCHO}_2} = -254.\underline{508} = -255 \text{ kJ/mol}$$

- 6.90 The equation is



Use the molar mass of 60.05 g/mol to convert -52.0 kJ/3.58 g to ΔH per mol of acid.

$$\frac{-52.0 \text{ kJ}}{3.58 \text{ g HC}_2\text{H}_3\text{O}_2} \times \frac{60.05 \text{ g HC}_2\text{H}_3\text{O}_2}{1 \text{ mol HC}_2\text{H}_3\text{O}_2} = -87\dot{2}.2 = -872 \text{ kJ/mol}$$

- 6.91 The heat gained by the water at the lower temperature equals the heat lost by the water at the higher temperature. Each heat term is mass \times sp ht $\times \Delta T$. This gives

$$(54.9 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(T_f - 31.5^\circ\text{C}) = (21.0 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(52.7^\circ\text{C} - T_f)$$

The specific heat and the units can be canceled from both sides to give

$$(54.9)(T_f - 31.5^\circ\text{C}) = (21.0)(52.7^\circ\text{C} - T_f)$$

After rearranging, you get $75.9 T_f = 2836.05$. This gives $T_f = 37.\dot{3}6 = 37.4^\circ\text{C}$.

- 6.92 The heat gained by the water at the lower temperature equals the heat lost by the water at the higher temperature. Each heat term is mass \times sp ht $\times \Delta T$. This gives

$$(45.4 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(T_f - 35.7^\circ\text{C}) = (20.5 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(66.2^\circ\text{C} - T_f)$$

The specific heat and the units can be canceled from both sides to give

$$(45.4)(T_f - 31.5^\circ\text{C}) = (20.5)(52.7^\circ\text{C} - T_f)$$

After rearranging, you get $65.9 T_f = 2977.88$. This gives $T_f = 45.\dot{1}8 = 45.2^\circ\text{C}$.

- 6.93 Divide the 235 J heat by the mass of lead and the Δt to obtain the specific heat.

$$\text{Specific heat} = \frac{235 \text{ J}}{121.6 \text{ g } (35.5^\circ\text{C} - 20.4^\circ\text{C})} = 0.12798 = \frac{0.128 \text{ J}}{\text{g} \cdot ^\circ\text{C}}$$

- 6.94 Divide the 47.0 J heat by the mass of copper and the Δt to obtain the specific heat.

$$\text{Specific heat} = \frac{47.0 \text{ J}}{35.4 \text{ g } (3.45^\circ\text{C})} = 0.38\dot{4}8 = \frac{0.385 \text{ J}}{\text{g} \cdot ^\circ\text{C}}$$

- 6.95 The energy used to heat the Zn comes from cooling the water. Calculate q for water:

$$q_{\text{wat}} = \text{specific heat} \times \text{mass} \times \Delta t$$

$$q_{\text{wat}} = \frac{4.18 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times 50.0 \text{ g} \times (96.68^\circ\text{C} - 100.00^\circ\text{C}) = -693.88 \text{ J}$$

The sign of q for the Zn is the reverse of the sign of q for water because the Zn is absorbing heat:

$$q_{\text{met}} = -(q_{\text{wat}}) = -(-693.88) = 693.88 \text{ J}$$

$$\text{Specific heat} = \frac{693.88 \text{ J}}{25.3 \text{ g} (96.68^\circ\text{C} - 25.00^\circ\text{C})} = 0.3826 = \frac{0.383 \text{ J}}{\text{g} \cdot ^\circ\text{C}}$$

- 6.96 The energy given up by the metal is used to heat the water. Calculate q for water:

$$q_{\text{wat}} = \text{specific heat} \times \text{mass} \times \Delta t$$

$$q_{\text{wat}} = \frac{4.18 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \times 26.7 \text{ g} \times (30.00^\circ\text{C} - 25.00^\circ\text{C}) = 558.03 \text{ J}$$

The sign of q for the metal is the reverse of the sign of q for water because the metal is giving up heat:

$$q_{\text{met}} = -(q_{\text{wat}}) = -558.03 \text{ J}$$

$$\text{Specific heat} = \frac{-558.03 \text{ J}}{19.6 \text{ g} (30.00^\circ\text{C} - 61.67^\circ\text{C})} = 0.8989 = \frac{0.899 \text{ J}}{\text{g} \cdot ^\circ\text{C}}$$

- 6.97 First, multiply the molarities by the volumes, in liters, to get the moles of NaOH and the moles of HCl, and thus determine the limiting reactant.

$$\text{Mol NaOH} = M \times V = 0.996 \text{ M} \times 0.0141 \text{ L} = 0.01404 \text{ mol}$$

$$\text{Mol HCl} = M \times V = 0.905 \text{ M} \times 0.0323 \text{ L} = 0.02923 \text{ mol}$$

Therefore, NaOH is the limiting reactant. The total volume of the system is 14.1 mL + 32.3 mL = 46.4 mL. Since the density of water is 1.00 g/mL, the total mass of the system is 46.4 g. Next, set the heat released by the reaction (mol \times enthalpy of reaction) equal to the heat absorbed by the water (m \times sp ht. \times ΔT). This gives

$$(0.01404 \text{ mol})(55.8 \times 10^3 \text{ J/mol}) = (46.4 \text{ g})(4.184 \text{ J/g} \cdot ^\circ\text{C})(T_f - 21.6^\circ\text{C})$$

After dividing, this gives $T_f - 21.6^\circ\text{C} = 4.035^\circ\text{C}$, or $T_f = 25.63 = 25.6^\circ\text{C}$

- 6.98 First, multiply the molarities by the volumes, in liters, to get the moles of KOH and the moles of HBr and thus determine the limiting reactant.

$$\text{Mol KOH} = M \times V = 1.05 \text{ M} \times 0.0291 \text{ L} = 0.030555 \text{ mol}$$

$$\text{Mol HBr} = M \times V = 1.07 \text{ M} \times 0.0209 \text{ L} = 0.022363 \text{ mol}$$

Therefore, HBr is the limiting reactant. The total volume of the system is 29.1 mL + 20.9 mL = 50.0 mL. Since the density of water is 1.00 g/mL, the total mass of the system is 50.0 g. Next, set the heat released by the reaction (mol x enthalpy of reaction) equal to the heat absorbed by the water (m x sp ht. x ΔT). This gives

$$(0.022363 \text{ mol})(55.8 \times 10^3 \text{ J/mol}) = (50.0 \text{ g})(4.184 \text{ J/g}\cdot^\circ\text{C})(T_f - 21.8^\circ\text{C})$$

After dividing, this gives $T_f - 21.8^\circ\text{C} = 5.964^\circ\text{C}$, or $T_f = 27.76 = 27.8^\circ\text{C}$

- 6.99 Use Δt and the heat capacity of 547 J/ $^\circ\text{C}$ to calculate q:

$$q = C\Delta t = (547 \text{ J}/^\circ\text{C})(36.66 - 25.00)^\circ\text{C} = 6.378 \times 10^3 \text{ J} (6.378 \text{ kJ})$$

Energy is released in the solution process in raising the temperature, so ΔH is negative:

$$\Delta H = \frac{-6.378 \text{ kJ}}{6.48 \text{ g LiOH}} \times \frac{23.95 \text{ g LiOH}}{1 \text{ mol LiOH}} = -23.57 = -23.6 \text{ kJ/mol}$$

- 6.100 Use Δt and the heat capacity of 682 J/ $^\circ\text{C}$ to calculate q:

$$q = C\Delta t = (682 \text{ J}/^\circ\text{C})(14.14 - 25.00)^\circ\text{C} = -7.4065 \times 10^3 \text{ J} (-7.4065 \text{ kJ})$$

Energy is absorbed in the solution process in lowering the temperature, so the sign of ΔH must be reversed, making the heat positive:

$$\Delta H = \frac{7.4065 \text{ kJ}}{21.45 \text{ g KNO}_3} \times \frac{101.1 \text{ g KNO}_3}{1 \text{ mol KNO}_3} = 34.91 = 34.9 \text{ kJ/mol}$$

- 6.101 Use Δt and the heat capacity of 13.43 kJ/ $^\circ\text{C}$ to calculate q:

$$q = C\Delta t = (13.43 \text{ kJ}/^\circ\text{C})(35.84 - 25.00)^\circ\text{C} = 145.581 \text{ kJ}$$

As in the previous two problems, the sign of ΔH must be reversed, making the heat negative:

$$\Delta H = \frac{-145.581 \text{ kJ}}{10.00 \text{ g HC}_2\text{H}_3\text{O}_2} \times \frac{60.05 \text{ g HC}_2\text{H}_3\text{O}_2}{1 \text{ mol HC}_2\text{H}_3\text{O}_2} = -874.21 = -874.2 \text{ kJ/mol}$$

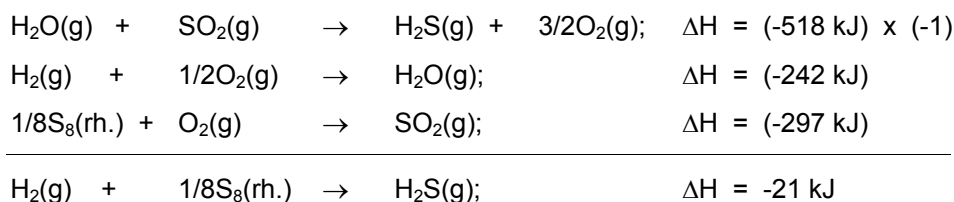
6.102 Use Δt and the heat capacity of $15.8 \text{ kJ/}^\circ\text{C}$ to calculate q :

$$q = C\Delta t = (15.8 \text{ kJ/}^\circ\text{C})(20.54 - 20.00)^\circ\text{C} = 8.53 \text{ kJ}$$

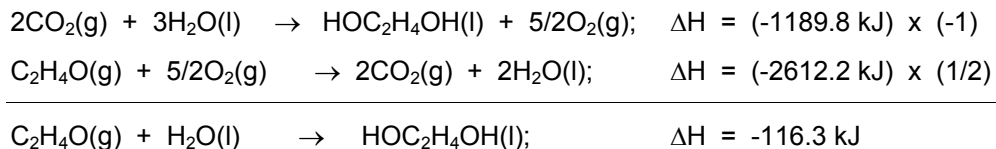
As in the previous three problems, the sign of ΔH must be reversed, making the heat negative:

$$\Delta H = \frac{-8.53 \text{ kJ}}{10.00 \text{ g sugar}} \times \frac{150.1 \text{ g sugar}}{1 \text{ mol sugar}} = -2.336 \times 10^3 = -2.3 \times 10^3 \text{ kJ/mol}$$

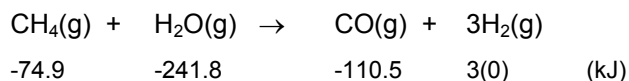
6.103 Using the equations in the data, reverse the direction of the first reaction, and reverse the sign of its ΔH . Then add the second and third equations and their ΔH 's.



6.104 Using the equations in the data, reverse the direction of the reaction involving oxidation of the glycol, and reverse the sign of its ΔH . Then add half of the reaction involving oxidation of $\text{C}_2\text{H}_4\text{O}$, and half of its ΔH .

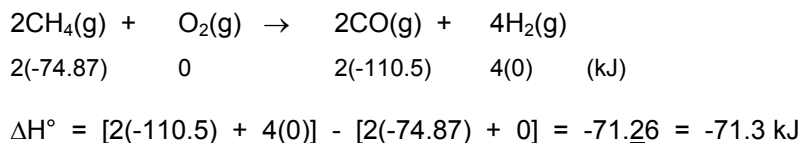


6.105 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation.

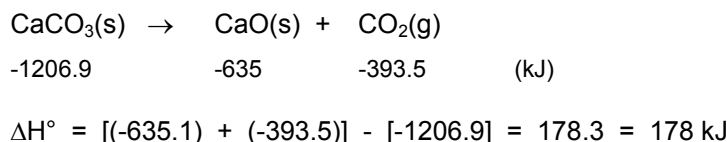


$$\Delta H^\circ = [-110.5 + 3(0)] - [(-74.87) + (-241.8)] = 206.17 = 206.2 \text{ kJ}$$

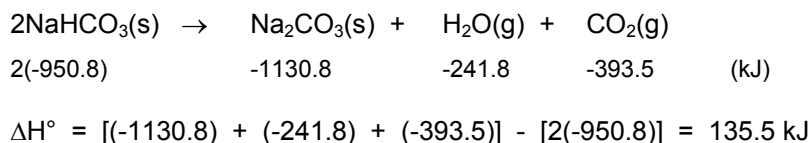
- 6.106 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation.



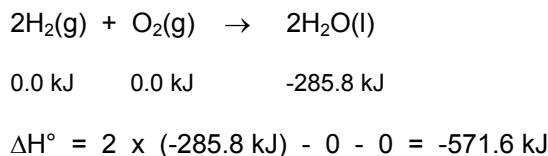
- 6.107 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation. The ΔH_f° of -635 kJ/mol CaO is given in the problem.



- 6.108 Write the ΔH° values (Table 6.2) underneath each compound in the balanced equation.



- 6.109 Calculate the molar heat of reaction from the equation with the ΔH_f° values below each substance; then convert to the heat for the reaction at 25°C .



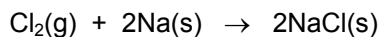
The moles of oxygen in 2.000 L with a density of 1.11 g/L is:

$$2.000 \text{ L O}_2 \times \frac{1.11 \text{ g O}_2}{1 \text{ L O}_2} \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} = 0.069\bar{37} \text{ mol O}_2$$

The heat of reaction from 0.06937 mol O_2 is:

$$0.69\bar{37} \text{ mol O}_2 \times \frac{-571.6 \text{ kJ}}{\text{mol O}_2} = -39.\underline{654} = -39.7 \text{ kJ}$$

- 6.110 Calculate the molar heat of reaction from the equation with the ΔH_f° values below each substance; then convert to the heat for the reaction at 25°C.



$$0.0 \text{ kJ} \quad 0.0 \text{ kJ} \quad -411.1 \text{ kJ}$$

$$\Delta H^\circ = 2 \times (-411.1 \text{ kJ}) - 0 - 0 = -822.2 \text{ kJ}$$

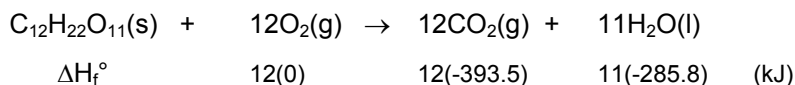
The moles of chlorine in 4.000 L Cl_2 with a density of 2.46 g/L is:

$$4.000 \text{ L Cl}_2 \times \frac{2.46 \text{ g Cl}_2}{\text{L Cl}_2} \times \frac{1 \text{ mol Cl}_2}{70.9 \text{ g Cl}_2} = 0.1387 \text{ mol Cl}_2$$

The heat of reaction from 0.1387 mol Cl_2 is:

$$0.1387 \text{ mol Cl}_2 \times \frac{-822.2 \text{ kJ}}{\text{mol Cl}_2} = -114.03 = -114 \text{ kJ}$$

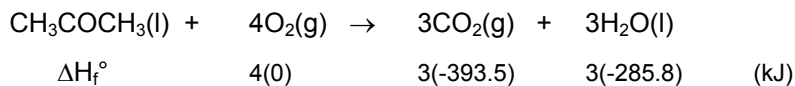
- 6.111 Let ΔH_f° = the unknown standard enthalpy of formation for sucrose. Then write this symbol and the other ΔH_f° values underneath each compound in the balanced equation, and solve for ΔH_f° using the ΔH° of -5641 kJ for the reaction.



$$\Delta H^\circ = -5641 = [12(-393.5) + 11(-285.8)] - [\Delta H_f^\circ + 12(0)]$$

$$\Delta H_f^\circ = -2225 \text{ kJ/mol sucrose}$$

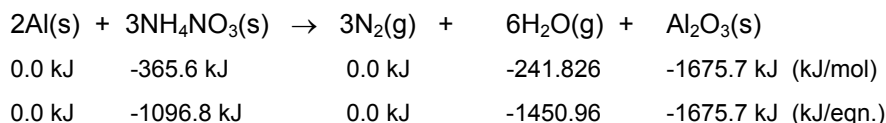
- 6.112 Let ΔH_f° = the unknown standard enthalpy of formation for acetone. Then write this symbol and the other ΔH_f° values underneath each compound in the balanced equation, and solve for ΔH_f° using the ΔH° of -1791 kJ for the reaction.



$$\Delta H^\circ = -1791 = [3(-393.5) + 3(-285.8)] - [\Delta H_f^\circ + 4(0)]$$

$$\Delta H_f^\circ = -246.9 = -247 \text{ kJ/mol acetone}$$

- 6.113 First, calculate the heat evolved from the molar amounts represented by the balanced equation. From Appendix C, obtain the individual heats of formation, and write those below the reactants and products in the balanced equation. Then, multiply the molar heats of formation by the number of moles in the balanced equation, and write those products below the molar heats of formation.



The total heat of reaction of two moles of Al and three moles of NH_4NO_3 is

$$\Delta H = -1450.8 - 1675.7 - (-1096.8 \text{ kJ}) = -2029.7 \text{ kJ}$$

Now, 245 kJ represents the following fraction of the total heat of reaction:

$$\frac{245 \text{ kJ}}{2029.7 \text{ kJ}} = 0.12074$$

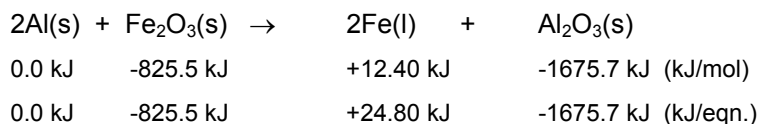
Thus, 245 kJ requires the fraction 0.12074 of the moles of each reactant: 0.12074×2 , or 0.24141, mol of Al and 0.12068×3 , or 0.36209, mol of NH_4NO_3 . The mass of each reactant and the mass of the mixture are as follows:

$$0.24141 \text{ mol Al} \times 26.98 \text{ g/mol Al} = 6.5132 \text{ g of Al}$$

$$0.36209 \text{ mol NH}_4\text{NO}_3 \times 80.04 \text{ g/mol NH}_4\text{NO}_3 = 28.9816 \text{ g NH}_4\text{NO}_3$$

$$6.5132 \text{ g Al} + 28.9816 \text{ g NH}_4\text{NO}_3 = 35.49 = 35.5 \text{ g of the mixture}$$

- 6.114 First, calculate the heat evolved from the molar amounts represented by the balanced equation. From Appendix C, obtain the individual heats of formation, and write those below the reactants and products in the balanced equation. Then, multiply the molar heats of formation by the number of moles in the balanced equation, and write those products below the molar heats of formation.



The total heat of reaction of two moles of Al and one mole of Fe_2O_3 is

$$\Delta H = +24.80 - 1675.7 - (-825.5 \text{ kJ}) = -825.40 \text{ kJ}$$

(continued)

Now, 348 kJ represents the following fraction of the total heat of reaction:

$$\frac{348 \text{ kJ}}{825.40 \text{ kJ}} = 0.4216$$

Thus, 348 kJ requires the fraction 0.4216 of the moles of each reactant: 0.4216×2 , or 0.8432, mol of Al and 0.4216×1 , or 0.4216, mol of Fe_2O_3 . The mass of each reactant and the mass of the mixture are as follows:

$$0.8432 \text{ mol Al} \times 26.98 \text{ g/mol Al} = 22.750 \text{ g Al}$$

$$0.4216 \text{ mol Fe}_2\text{O}_3 \times 159.7 \text{ g/mol Fe}_2\text{O}_3 = 67.331 \text{ g Fe}_2\text{O}_3$$

$$22.750 \text{ g Al} + 67.331 \text{ g Fe}_2\text{O}_3 = 90.081 = 90.1 \text{ g of the mixture}$$

■ Solutions to Cumulative-Skills Problems

- 6.115 The heat lost, q , by the water(s) with a temperature higher than the final temperature must be equal to the heat gained, q , by the water(s) with a temperature lower than the final temperature.

$$q(\text{lost by water at higher temp}) = q(\text{gained by water at lower temp})$$

$$\Sigma (s \times m \times \Delta t) = \Sigma (s \times m \times \Delta t)$$

Divide both sides of the equation by the specific heat, s , to eliminate this term, and substitute the other values. Since the mass of the water at 50.0°C is greater than the sum of the masses of the other two waters, assume the final temperature will be greater than 37°C and greater than 15°C . Use this to set up the three Δt expressions, and write one equation in one unknown, letting t equal the final temperature. Simplify by omitting the “grams” from 45.0 g, 25.0 g, and 15.0 g.

$$\Sigma (m \times \Delta t) = \Sigma (m \times \Delta t)$$

$$45.0 \times (50.0^\circ\text{C} - t) = 25.0 \times (t - 15.0^\circ\text{C}) + 15.0 \times (t - 37^\circ\text{C})$$

$$2250^\circ\text{C} - 45.0 t = 25.0 t - 375^\circ\text{C} + 15.0 t - 555.0^\circ\text{C}$$

$$-85.0 t = -3180^\circ\text{C}$$

$$t = 37.41 = 37.4^\circ\text{C} \quad (\text{the final temperature})$$

- 6.116 The heat lost, q , by the substances(s) with a temperature higher than the final temperature must be equal to the heat gained, q , by the substances(s) with a temperature lower than the final temperature.

$$q(\text{lost}) = q(\text{gained})$$

$$\Sigma (s \times m \times \Delta t) = \Sigma (s \times m \times \Delta t)$$

Since the temperature of the iron, 95.0°C is far larger than the temperatures of the other two substances, assume the final temperature will be greater than 35.5°C and greater than 25°C . Then, write the heat term for the iron on the left and the sum of the heat terms for water and ethanol on the right. Since the masses, m , are equal, divide both sides by m to eliminate this term, and solve one equation in one unknown, letting t equal the final temperature. (To simplify the setup, only the temperature unit is retained.)

$$\Sigma (s \times \Delta t) = \Sigma (s \times \Delta t)$$

$$0.449 \times (95.0^\circ\text{C} - t) = 2.43 \times (t - 35.5^\circ\text{C}) + 4.18 \times (t - 25.0^\circ\text{C})$$

$$42.655^\circ\text{C} - 0.449 t = 2.43 t - 86.265^\circ\text{C} + 4.18 t - 104.5^\circ\text{C}$$

$$-7.059 t = -233.42^\circ\text{C}$$

$$t = 33.\underline{06} = 33.1^\circ\text{C} \text{ (the final temperature)}$$

The final temperature is actually less than that of the ethanol, but rewriting the equation to account for this still gives the same final temperature.

- 6.117 First, calculate the mole fraction of each gas in the product, assuming 100 g product:

$$\text{Mol CO} = 33 \text{ g CO} \times 1 \text{ mol CO}/28.01 \text{ g CO} = 1.\underline{178} \text{ mol CO}$$

$$\text{Mol CO}_2 = 67 \text{ g CO}_2 \times 1 \text{ mol CO}_2/44.01 \text{ g CO}_2 = 1.\underline{522} \text{ mol CO}_2$$

$$\text{Mol frac. CO} = \frac{1.178 \text{ mol CO}}{(1.178 + 1.522) \text{ mol}} = 0.43\underline{63}$$

$$\text{Mol frac. CO}_2 = \frac{1.522 \text{ mol CO}_2}{(1.178 + 1.522) \text{ mol}} = 0.56\underline{37}$$

Now, calculate the starting moles of C, which equal the total moles of CO and CO₂:

$$\begin{aligned} \text{Starting mol C(s)} &= 1.00 \text{ g C} \times 1 \text{ mol C}/12.01 \text{ g C} \\ &= 0.083\underline{26} \text{ mol C} = \text{mol CO} + \text{CO}_2 \end{aligned}$$

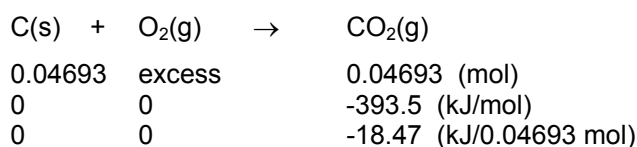
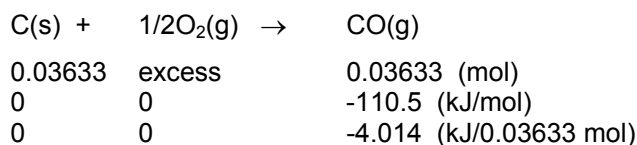
(continued)

Use the mol fractions to convert mol CO + CO₂ to mol CO and mol CO₂:

$$\text{Mol CO} = \frac{0.4363 \text{ mol CO}}{1 \text{ mol total}} \times 0.08326 \text{ mol total} = 0.03633 \text{ mol CO}$$

$$\text{Mol CO}_2 = \frac{0.5637 \text{ mol CO}_2}{1 \text{ mol total}} \times 0.08326 \text{ mol total} = 0.04693 \text{ mol CO}_2$$

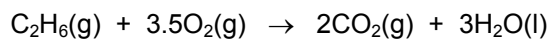
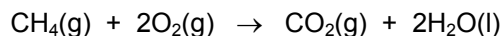
Now, use enthalpies of formation (Table 6.2) to calculate the heat of combustion for both:



$$\text{Total } \Delta H = -4.014 + (-18.47) = -22.48 = -22 \text{ kJ}$$

$$\text{Heat released} = 22 \text{ kJ}$$

6.118 The balanced equations are



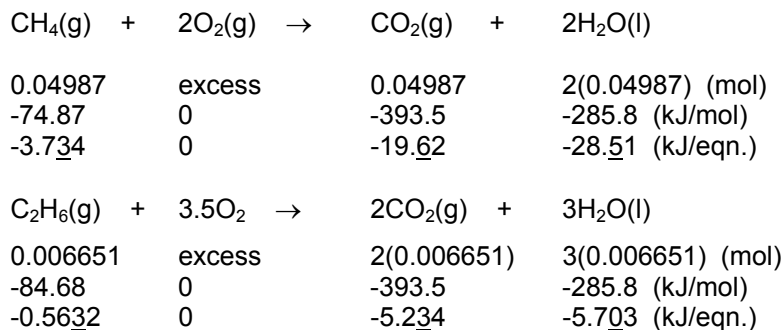
First, calculate the moles of CH₄ and C₂H₆ (let 80.0% = 0.800 g, and 20.0% = 0.200 g).

$$\text{Mol CH}_4 = 0.800 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} = 0.04987 \text{ mol CH}_4$$

$$\text{Mol C}_2\text{H}_6 = 0.200 \text{ g C}_2\text{H}_6 \times \frac{1 \text{ mol C}_2\text{H}_6}{30.07 \text{ g C}_2\text{H}_6} = 0.006651 \text{ mol C}_2\text{H}_6$$

(continued)

Now, use the enthalpies of formation (Table 6.2) to calculate the total heat:



For the combustion of CH_4 :

$$\Delta H = [(-28.51) + (-19.62)] - [(-3.734) + 0] = -44.395 \text{ kJ}$$

For the combustion of C_2H_6 :

$$\Delta H = [(-5.234) + (-5.703)] - [(-0.5632) + 0] = -10.373 \text{ kJ}$$

For the combustion of the total (1.00 g) of CH_4 and C_2H_6 :

$$\Delta H = -44.395 + (-10.373) = -54.768 = -54.8 \text{ kJ}$$

The heat evolved = 54.8 kJ

- 6.119 The ΔH° for the reactions to produce CO and CO_2 in each case is equal to the ΔH_f° of the respective gases. Thus ΔH° to produce CO is -110.5 kJ/mol, and ΔH° to produce CO_2 is -393.5 kJ/mol. Since one mol of graphite is needed to produce one mol of either CO or CO_2 , one equation in one unknown can be written for the heat produced using m for the moles of CO and $(2.00 - m)$ for moles of CO_2 .

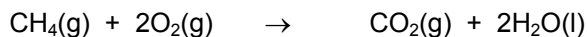
$$-481 \text{ kJ} = m \times (-110.5 \text{ kJ}) + (2.00 - m) \times (-393.5 \text{ kJ})$$

$$-481 = -110.5 m - 787 + 393.5 m$$

$$\text{mol CO} = m = 1.081; \text{ mass CO} = 1.081 \times 28 \text{ g/mol} = 30.26 = 30.3 \text{ g}$$

$$\begin{aligned} \text{mol CO}_2 &= (2.00 - m) = 0.919; \text{ mass CO}_2 = 0.919 \text{ mol} \times 44.0 \text{ g/mol} \\ &= 40.4 = 40. \text{ g} \end{aligned}$$

6.120 The ΔH° values for each reaction are as follows:



$$\Delta H^\circ = -393.5 \text{ kJ} + 2(-285.8 \text{ kJ}) - (-74.87 \text{ kJ}) - 0 = -890.23 \text{ kJ}$$



$$\Delta H^\circ = 2(-393.5 \text{ kJ}) + 2(-285.8 \text{ kJ}) - (52.47 \text{ kJ}) - 0 = -1411.07 \text{ kJ}$$

Since 10.0 g of CH_4 and C_2H_4 produce 520 kJ of heat, one equation in one unknown can be written for the heat produced using m for the mass of CH_4 and $(10.00 - m)$ for the mass of C_2H_4 . Use $m/16.0$ for the moles of CH_4 and $(10.00 - m)/28.0$ for the moles of C_2H_4 .

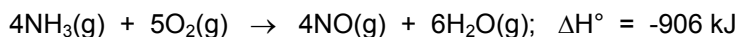
$$-520 \text{ kJ} = (m/16.0) \times (-890.23 \text{ kJ}) + [(10.00 - m)/28.0] \times (-1411.07 \text{ kJ})$$

$$-520 = -55.63 m + (-503.95) + 50.39 m$$

$$\text{mass CH}_4 = m = 16.05 \div 5.24 = 3.062 = 3.1 \text{ g}$$

$$\text{mass percentage CH}_4 = (3.06 \text{ g} \div 10.0 \text{ g}) \times 100\% = 3.06 = 3.1\%$$

6.121 The equation is



First, determine the limiting reactant by calculating the moles of NH_3 and of O_2 ; then, assuming one of the reactants is totally consumed, calculate the moles of the other reactant needed for the reaction.

$$10.0 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} = 0.5872 \text{ mol NH}_3$$

$$20.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.6250 \text{ mol O}_2$$

$$0.6250 \text{ mol O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.500 \text{ mol NH}_3 \text{ needed}$$

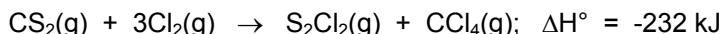
Because NH_3 is present in excess of what is needed, O_2 must be the limiting reactant. Now calculate the heat released on the basis of the complete reaction of 0.622250 mol O_2 :

(continued)

$$\Delta H = \frac{-906 \text{ kJ}}{5 \text{ mol O}_2} \times 0.6250 \text{ mol O}_2 = -113.25 = -113 \text{ kJ}$$

The heat released by the complete reaction of the 20.0 g (0.6250 mol) of O₂(g) is 113 kJ.

6.122 The equation is



First, determine the limiting reactant by calculating the moles of CS₂ and of Cl₂; then, assuming one of the reactants is totally consumed, calculate the moles of the other reactant needed for the reaction.

$$10.0 \text{ g CS}_2 \times \frac{1 \text{ mol CS}_2}{76.13 \text{ g CS}_2} = 0.1314 \text{ mol CS}_2$$

$$10.0 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.91 \text{ g Cl}_2} = 0.1410 \text{ mol Cl}_2$$

$$0.1410 \text{ mol Cl}_2 \times \frac{1 \text{ mol CS}_2}{3 \text{ mol Cl}_2} = 0.04700 \text{ mol CS}_2$$

Because CS₂ is present in excess of what is needed, Cl₂ must be the limiting reactant. Now, calculate the heat released on the basis of the complete reaction of 0.1410 mol Cl₂:

$$\Delta H = \frac{-230 \text{ kJ}}{3 \text{ mol Cl}_2} \times 0.1410 \text{ mol Cl}_2 = -10.810 = -10.8 \text{ kJ}$$

The heat released by the complete reaction of 10.0 g of Cl₂(g) is 10.8 kJ.

6.123 The equation is



- a. To find the heat evolved from the production of 1.00 L of NH₃, convert the 1.00 L to mol NH₃ using the density and molar mass (17.03 g/mol). Then convert the moles to heat (ΔH) using ΔH°:

(continued)

$$\text{moles NH}_3 = \frac{d \times V}{\text{Mwt}} = \frac{0.696 \text{ g/L} \times 1 \text{ L}}{17.03 \text{ g/mol}} = 0.04087 \text{ mol NH}_3$$

$$\Delta H = \frac{-91.8 \text{ kJ}}{2 \text{ mol NH}_3} \times 0.04087 \text{ mol NH}_3 = -1.876 \text{ (1.88 kJ heat evolved)}$$

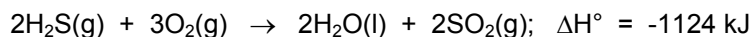
- b. First, find the moles of N₂ using the density and molar mass (28.02 g/mol). Then convert to the heat needed to raise the N₂ from 25°C to 400°C:

$$\text{moles N}_2 = \frac{d \times V}{\text{Mwt}} = \frac{1.145 \text{ g/L} \times 0.500 \text{ L}}{28.02 \text{ g/mol}} = 0.02043 \text{ mol N}_2$$

$$0.02043 \text{ mol N}_2 \times \frac{29.12 \text{ J}}{\text{mol} \cdot ^\circ\text{C}} \times (400 - 25)^\circ\text{C} = 223.2 \text{ J (0.2232 kJ)}$$

$$\text{Percent heat for N}_2 = \frac{0.2232 \text{ kJ}}{1.876 \text{ kJ}} \times 100\% = 11.89 = 11.9\%$$

6.124 The equation is



- a. First, calculate the moles of SO₂ using the density and molar mass (64.07 g/mol). Then convert the moles to heat (ΔH) using ΔH°:

$$\text{moles SO}_2 = \frac{d \times V}{\text{Mwt}} = \frac{2.62 \text{ g/L} \times 1 \text{ L}}{64.07 \text{ g/mol}} = 0.04089 \text{ mol SO}_2$$

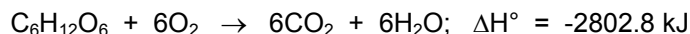
$$\Delta H = \frac{-1124 \text{ kJ}}{2 \text{ mol SO}_2} \times 0.04089 \text{ mol SO}_2 = -22.98 \text{ (23.0 kJ heat evolved)}$$

- b. Recall that 1.000 L of SO₂ = 0.04089 mol SO₂. Use this with the molar heat capacity of 30.2 J/(mol·°C) to calculate the heat needed to raise the temperature of SO₂ from 25°C to 500°C:

$$0.04089 \text{ mol SO}_2 \times \frac{30.2 \text{ J}}{(\text{mol} \cdot ^\circ\text{C})} \times (500 - 25)^\circ\text{C} = 586.6 \text{ J (0.5866 kJ)}$$

$$\text{Percent heat for SO}_2 = \frac{0.5866 \text{ kJ}}{22.98 \text{ kJ}} \times 100\% = 2.552 = 2.55\%$$

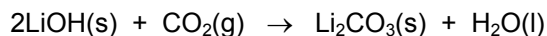
6.125 The glucose equation is



Convert the 2.50×10^3 kcal to mol of glucose using the ΔH° of -2802.8 kJ for the reaction and the conversion factor of 4.184 kJ/kcal:

$$2.50 \times 10^3 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1.000 \text{ kcal}} \times \frac{1 \text{ mol glucose}}{2802.8 \text{ kJ}} = 3.7319 \text{ mol glucose}$$

Next, convert mol glucose to mol LiOH using the above equation for glucose and the equation for LiOH:

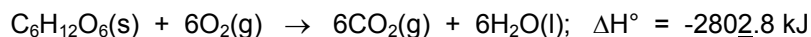


$$3.7319 \text{ mol glucose} \times \frac{6 \text{ mol CO}_2}{1 \text{ mol glucose}} \times \frac{2 \text{ mol LiOH}}{1 \text{ mol CO}_2} = 44.783 \text{ mol LiOH}$$

Finally, use the molar mass of LiOH to convert moles to mass:

$$44.783 \text{ mol LiOH} \times \frac{23.95 \text{ g LiOH}}{1 \text{ mol LiOH}} = 1.0725 \times 10^3 \text{ g (1.07 kg) LiOH}$$

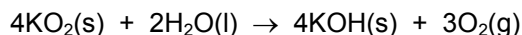
6.126 Using $\Delta H_f^\circ = -1273 \text{ kJ}$ for glucose, the conversion of glucose to heat and ΔH° in the body is



Convert the 1.00×10^2 kcal to mol of glucose, using the ΔH° of -2802.8 kJ:

$$100 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1.000 \text{ kcal}} \times \frac{1 \text{ mol glucose}}{2802.8 \text{ kJ}} = 0.14927 \text{ mol glucose}$$

Next, convert mol glucose to mol KO₂ using the above equation for glucose and the equation for KO₂:



$$0.14927 \text{ mol glucose} \times \frac{6 \text{ mol O}_2}{1 \text{ mol glucose}} \times \frac{4 \text{ mol KO}_2}{3 \text{ mol O}_2} = 1.1942 \text{ mol KO}_2$$

$$1.1942 \text{ mol KO}_2 \times \frac{71.1 \text{ g KO}_2}{1 \text{ mol KO}_2} = 84.91 = 84.9 \text{ g KO}_2$$

Filename: chapter6.doc
Directory: D:\hmco\chemistry\general\ebbing\general_chem\8e\instructors\
solutions_manual
Template: C:\Documents and Settings\willsoj\Application
Data\Microsoft\Templates\Normal.dot
Title: 6
Subject:
Author: David Bookin
Keywords:
Comments:
Creation Date: 10/22/2000 11:01 PM
Change Number: 73
Last Saved On: 11/21/2003 3:41 PM
Last Saved By: David Bookin
Total Editing Time: 763 Minutes
Last Printed On: 1/9/2004 3:52 PM
As of Last Complete Printing
Number of Pages: 37
Number of Words: 8,417 (approx.)
Number of Characters: 47,980 (approx.)