Chapter 8 Practice - Key

8.1 Energy, Work, and Heat

Energy Equations and Constants
$\Delta E = q + w$
$q = ms\Delta T$
$q = C\Delta T$
s_{water} = 1 cal/g·°C = 4.184 J/g·°C

1. Convert 3,415.5 kilocalories to kilojoules.

$$3415.5\ \cancel{\text{kcal}} \times \frac{4.184\ \text{kJ}}{1\ \cancel{\text{kcal}}} = 14{,}290\ \text{kJ}$$

2. Convert 95.2 kWh to J.

$$95.2\ \cancel{\text{kWh}} \times \frac{3.6\times10^6\ \text{J}}{1\ \cancel{\text{kWh}}} = 3.43\times10^8\ \text{J}$$

3. As a sample of natural gas burns, it releases 55.0 J of heat and performs 23.0 J of work on its surroundings. What is the total change in energy for the chemical system?

$$\Delta E = q + w = (-55.0\ \text{J}) + (-23.0\ \text{J}) = -78.0\ \text{J}$$

Common Units of Energy
1 joule = 1 kg·m^2/s^2
1 calorie = 4.184 J
1,000 calories = 1 kcal = 1 Calorie
1 British Thermal Unit (BTU) = 1,055 J
1 kilowatt-hour (kWh) = 3.6 × 10^6 J

8.2 Heat Energy and Temperature

4. An engineer tests the thermal properties of a metal alloy. Using a 50.0-gram sample, she finds that adding 485 J of heat energy to the alloy causes a temperature change of 4.10 °C. What is the specific heat of this alloy?

$$s = \frac{q}{m\Delta T} = \frac{485\ \text{J}}{(50.0\ \text{g})\ (4.10\ ^\circ\text{C})} = 2.37\ \text{J/g}\cdot{}^\circ\text{C}$$

5. Water has a specific heat of 4.184 J/g·°C. If 300.0 grams of water absorb 2.50 × 10^4 J of energy, how much will the temperature change?

$$\Delta T = \frac{q}{ms} = \frac{2.50\times10^4\ \cancel{\text{J}}}{(300.0\ \cancel{\text{g}})(4.184\ \cancel{\text{J}}/\cancel{\text{g}}\cdot{}^\circ\text{C})} = 19.1\ ^\circ\text{C}$$

6. How much heat is required to raise 900.0 g of water by a temperature of 15.0 °C?

$$q = ms\Delta T = (900.0\ \text{g})\left(4.184\frac{\text{J}}{\cancel{\text{g}}}\cdot\cancel{^\circ\text{C}}\right)(15.0\ \cancel{^\circ\text{C}}) = 56{,}500\ \text{J}$$

7. An industrial reaction vessel is found to undergo a change in temperature of 0.061 °C for each kilojoule of energy absorbed. What is the heat capacity for this vessel in kJ/°C?

$$C = \frac{q}{\Delta T} = \frac{1\ \text{kJ}}{0.061\ ^\circ\text{C}} = 16.4\ \text{kJ/}^\circ\text{C}$$

8. In a calorimetry experiment, a reaction takes place in 90.0 g of an aqueous solution. The solution temperature drops from 25.8 °C to 17.5 °C. Calculate $q_{reaction}$ and $q_{solution}$ for this experiment. Assume the specific heat of the solution is identical to that of pure water.

$$q_{solution} = ms\Delta T = (90.0\ \text{g})\left(4.184\frac{\text{J}}{\cancel{\text{g}}}\cdot\cancel{^\circ\text{C}}\right)(-8.3\ \cancel{^\circ\text{C}}) = -3{,}100\ \text{J}$$

$$q_{\text{reaction}} = -q_{\text{solution}} = +3{,}100 \text{ J}$$

9. A chemist conducts a calorimetry experiment on an unknown metal. The mass of the sample is 24.4 g. The chemist heats the metal to 100 °C, then places it in a coffee-cup calorimeter containing 108.5 g of water. The temperature of the water rises from 21.3 to 25.0 °C.
   a. How much heat was absorbed by the water?

$$q_{\text{water}} = ms\Delta T = (108.5 \text{ g})\left(4.184 \frac{\text{J}}{\cancel{\text{g}}} \cdot \cancel{°\text{C}}\right)(3.7 \ \cancel{°\text{C}}) = +1{,}680 \text{ J}$$

(For the final answer, we would round to 2 significant digits. We have kept additional digits to carry through the calculation.)

   b. How much heat was released by the metal?

$$q_{\text{metal}} = -q_{\text{solution}} = -1{,}680 \text{ J}$$

   c. What is the specific heat of the metal?

$$\Delta T_{\text{metal}} = 25.0 \text{ °C} - 100.0 \text{ °C} = -75.0 \text{ °C}$$

$$s = \frac{q}{m\Delta T} = \frac{-1{,}680 \text{ J}}{(24.4 \text{ g})(-75.0 \text{ °C})} = 0.92 \text{ J/g} \cdot \text{°C}$$

10. A 1.0-gram sample of a candy bar was placed in a bomb calorimeter with a heat capacity of 3.54 kcal/°C. The sample was completely burned, causing the temperature of the calorimeter to rise by 1.36 °C. How many kilocalories of energy were stored in the candy bar?

$$q = C\Delta T = \left(3.54 \frac{\text{kcal}}{\cancel{°\text{C}}}\right)(1.36 \ \cancel{°\text{C}}) = 4.81 \text{ kcal}$$

8.3 Heat Energy and Chemical Reactions

11. A synthetic fuel blend is tested by bomb calorimetry. It is found that an 80.10-g sample releases 28,096 kJ of heat. What is the fuel value of this blend in kJ/g?

$$\text{fuel value} = \frac{28{,}096 \text{ kJ}}{80.10 \text{ g}} = 350.8 \text{ kJ/g}$$

12. Using the reaction enthalpy data shown, determine the amount of heat released in each situation:

$$2 \text{ Ca } (s) + O_2 \ (g) \rightarrow 2 \text{ CaO } (s) \qquad \Delta H_{\text{rxn}} = -1{,}269.8 \text{ kJ}$$

   a. If calcium reacts with 3.00 moles of O_2
   b. If 8.00 moles of calcium react with O_2
   c. If calcium and oxygen react to form 1.00 mole of CaO

$$\text{a. } 3.00 \ \cancel{\text{mol } O_2} \times \frac{-1{,}269.8 \text{ kJ}}{1 \ \cancel{\text{mol } O_2}} = -3{,}810 \text{ kJ}$$

$$\text{b. } 8.00 \ \cancel{\text{mol Ca}} \times \frac{-1{,}269.8 \text{ kJ}}{2 \ \cancel{\text{mol Ca}}} = -5{,}080 \text{ kJ}$$

$$\text{c. } 1.00 \ \cancel{\text{mol CaO}} \times \frac{-1{,}269.8 \text{ kJ}}{2 \ \cancel{\text{mol CaO}}} = -635 \text{ kJ}$$

13. Butane, C_4H_{10}, is the fuel used in many handheld lighters (see Figure 8.17). Based on the reaction enthalpy below, how many grams of butane are needed to produce 1,000 kJ of heat by this reaction?

$$2\ C_4H_{10}\ (g) + 13\ O_2\ (g) \rightarrow 8\ CO_2\ (g) + 10\ H_2O\ (l) \qquad \Delta H_{rxn} = -5{,}755\ \text{kJ}$$

$$-1{,}000\ \cancel{\text{kJ}} \times \frac{2\ \cancel{\text{mol}\ C_4H_{10}}}{-5{,}755\ \cancel{\text{kJ}}} \times \frac{58.12\ \text{g}\ C_4H_{10}}{1\ \cancel{\text{mol}\ C_4H_{10}}} = 20.20\ \text{g}\ C_4H_{10}$$