6.2: Energy and Properties (Exercises)

These are homework exercises to accompany Chapter 6 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Solutions are available below the questions.

Questions

6.1: Heat Flow

(click here for solutions)

Q6.1.1

Define potential energy and chemical potential energy.

Q6.1.2

What is one potential use for substances that have a large amount of chemical potential energy?

Q6.1.3

Describe what happens when two objects that have different temperatures come into contact with one another.

Q6.1.4

Distinguish between system and surroundings.

Q6.1.5

Distinguish between endothermic and exothermic.

Q6.1.6

Two different reactions are performed in two identical test tubes. In reaction A, the test tube becomes very warm as the reaction occurs. In reaction B, the test tube becomes cold. Which reaction is endothermic and which is exothermic? Explain.

Q6.1.7

What is the sign of q for an endothermic process? For an exothermic process?

Q6.1.8

Classify the following as endothermic or exothermic processes.

a. Boiling water
b. Sweating
c. Burning paper
d. Water freezing

Q6.1.9

Convert each value to the indicated units.

a. 150. kcal to Cal
b. 355 J to cal
c. 200. Cal to J
d. 225 kcal to cal
e. 3450. cal to kcal
f. 450. Cal to kJ
g. 175 kJ to cal

Q6.1.10

Equal amounts of heat are applied to 10.0 g samples of iron and aluminum, both originally at 25°C. Which one will be at the higher temperature?

Q6.1.11

Which sample will require more heat to increase the temperature by 10°C?

a. 25.0 g copper
b. 25.0 g lead

Q6.1.12

How much energy is required to heat 50.0 g of silver from 30°C to 50°C?

Q6.1.13

What is the final temperature when 125 J is applied to 20.0 g of lead, initially at 15°C?

Q6.1.14

How much energy is required to raise the temperature of 13.7 g of aluminum from 25.2°C to 61.9°C?

Q6.1.15

A 274 g sample of air is heated with 2250 J of heat and its temperature rises by 8.11°C. What is the specific heat of air at these conditions?

Q6.1.16

98.3 J of heat is supplied to 12.28 g of a substance, and its temperature rises by 5.42°C. What is the specific heat of the substance?

Q6.1.17

A quantity of ethanol is cooled from 47.9°C to 12.3°C and releases 3.12 kJ of heat. What is the mass of the ethanol sample?

Answers

6.1: Heat Flow

Q6.1.1

Potential energy is usually described as the energy of position. Chemical potential energy is energy stored within the chemical bonds of a substance.

Q6.1.2

Answers will vary. The most common example in every day life is the burning of fossil fuels to generate electricity or to run a vehicle.

Q6.1.3

The temperature of the hot object decreases and the temperature of the cold object increases as heat is transferred from the hot object to the cold object. The change in temperature of each depends on the identity and properties of each substance.

Q6.1.4

The system is the specific portion of matter being observed in an experiment and is designated by the experimenter. The surroundings is everything that is not the system.

Q6.1.5

Endothermic processes result in the gain of heat to the system while exothermic processes are associated with the loss of heat from the system.

Q6.1.6

Reaction A is exothermic because heat is leaving the system making the test tube feel hot. Reaction B is endothermic because heat is being absorbed by the system making the test tube feel cold.

Q6.1.7

q is positive for endothermic processes and q is negative for exothermic processes.

Q6.1.8

Classify the following as endothermic or exothermic processes.

a. Endothermic because heat is being added to the water to get it from the liquid state to the gas state.
b. Endothermic because energy is consumed to evaporate the moisture on your skin which lowers your temperature.
c. Exothermic because burning (also known as combustion) releases heat.
d. Exothermic because energy is exiting the system in order to go from liquid to solid. Another way to look at it is to consider the opposite process of melting. Energy is consumed (endothermic) to melt ice (solid to liquid) so the opposite process (liquid to solid) must be exothermic.

Q6.1.9

Convert each value to the indicated units.

a. $150\ kcal\left(\frac{1\ Cal}{1\ kcal}\right) = 150\ Cal$
b. $355\ J\left(\frac{1\ cal}{4.184\ J}\right) = 84.8\ cal$
c. $200.\ Cal\left(\frac{1000\ cal}{1\ Cal}\right)\left(\frac{4.184\ J}{1\ cal}\right) = 8.37 \times 10^5\ J$
d. $225\ kcal\left(\frac{1000\ cal}{1\ kcal}\right) = 2.25 \times 10^5\ cal$
e. $3450.\ cal\left(\frac{1 kcal\ kJ}{1000\ cal}\right) = 3.450\ kcal$
f. $450.\ Cal\left(\frac{1000\ cal}{1\ Cal}\right)\left(\frac{4.184\ J}{1\ cal}\right)\left(\frac{1\ kJ}{1000\ J}\right) = 1.88 \times 10^3\ kJ$ or $450.\ Cal\left(\frac{4.184\ kJ}{1\ Cal}\right) = 1.88 \times 10^3\ kJ$
g. $175\ kJ\left(\frac{1000\ J}{1\ kJ}\right)\left(\frac{1\ cal}{4.184\ J}\right) = 4.18 \times 10^4\ cal$

Q6.1.10

Iron has a specific heat capacity of 0.449 $\mathrm{J/g \cdot ^o C}$ which means it takes 0.449 J of energy to raise 1 gram of iron by 1$^{\text{o}}$C. Aluminum has a specific heat capacity of 0.897 $\mathrm{J/g \cdot ^o C}$ which means it takes 0.897 J of energy to raise 1 gram of aluminum by 1 $^{\text{o}}$C. When equal amounts of heat are applied, the temperature of the iron will increase more because it takes less energy (heat) to raise its temperature so iron will be at a higher temperature since they both start at 25$^{\text{o}}$C.

Q6.1.11

Both samples are the same mass so a comparison of the specific heat must be compared. Copper has a specific heat of 0.385 $\mathrm{J/g \cdot ^o C}$ which means it takes 0.385 J of energy to raise 1 gram of copper by 1^\text{o} \text{C}\). Lead has a specific heat of 0.129 $\mathrm{J/g \cdot ^o C}$ which means it takes 0.129 J of energy to raise 1 gram of copper by 1^\text{o} \text{C}\). More energy is needed to raise the temperature of copper so more heat will be needed to increase the temperature of copper by 10 ^\text{o} \text{C}\).

Q6.1.12

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ q &= 50.0\ g \cdot 0.233\ \frac{J}{g \cdot ^o\mathrm{C}} \cdot (50^o\mathrm{C} - 30^o\mathrm{C}) \\ q &= 50.0\ g \cdot 0.233\ \frac{J}{g \cdot ^o\mathrm{C}} \cdot (20^o\mathrm{C}) \\ q &= 233\ J \end{aligned}$$

Q6.1.13

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ 125\ J &= 20.0\ g \cdot 0.129\ \frac{J}{g \cdot ^o\mathrm{C}} \cdot (T_f - 15^o\mathrm{C}) \\ 48.4^o\mathrm{C} &= T_f - 15^o\mathrm{C} \\ T_f &= 63^o\mathrm{C} \end{aligned}$$

Q6.1.14

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ q &= 13.7\ g \cdot 0.897\ \frac{J}{g \cdot ^o\mathrm{C}} \cdot (61.9^o\mathrm{C} - 25.2^o\mathrm{C}) \\ q &= 13.7\ g \cdot 0.897\ \frac{J}{g \cdot ^o\mathrm{C}} \cdot (36.7^o\mathrm{C}) \\ q &= 451\ J \end{aligned}$$

Q6.1.15

$$
\begin{aligned}
q &= m \cdot C_p \cdot \Delta T \\
2250\ J &= 274\ g \cdot C_p\ \cdot 8.11^{\circ}\mathrm{C} \\
C_p &= 1.01 \frac{J}{g \cdot {}^{\circ}\mathrm{C}}
\end{aligned}
$$

Q6.1.16

$$
\begin{aligned}
q &= m \cdot C_p \cdot \Delta T \\
98.3\ J &= 12.28\ g \cdot C_p\ \cdot 5.42^{\circ}\mathrm{C} \\
C_p &= 1.48 \frac{J}{g \cdot {}^{\circ}\mathrm{C}}
\end{aligned}
$$

Q6.1.17

$$
\begin{aligned}
q &= m \cdot C_p \cdot \Delta T \\
-3.12\ kJ &= m \cdot 2.44 \frac{J}{g \cdot {}^{\circ}\mathrm{C}} (12.3^{\circ}\mathrm{C} - 47.9^{\circ}\mathrm{C}) \\
-3.12 \times 10^{3}\ J &= m \cdot 2.44 \frac{J}{g \cdot {}^{\circ}\mathrm{C}} (-35.6^{\circ}\mathrm{C}) \\
m &= 35.9\ g
\end{aligned}
$$
