CHAPTER 10

Energy

INTRODUCTION

Chemistry is about matter and its changes. As a part of all chemical reactions, energy (in the form of heat) is either released or required. Even though energy is converted from one form to another as a chemical reaction proceeds, the total amount of energy in the universe is constant (this is known as the law of conservation of energy). In this chapter you will learn about the nature of energy and energy sources, discover how we can predict whether or not a reaction will be spontaneous, and calculate the amount of energy needed to heat water and other substances.

CHAPTER DISCUSSION

Energy is defined as the ability to do work or produce heat. For example, imagine that we move a chair across a room. We do work on the chair because we need to exert a force to get the chair to move, and we move it over a certain distance (work is defined as a force acting over a distance). The heavier the chair, the more force required, so the more work is done (or the more energy is required). The farther we move the chair (the larger the distance), the more work is done, and the more energy is required. Heat is flow of energy due to a temperature difference. Although we often talk of heat as though it were a substance ("Close the window, you are letting the heat out"), heat is not a "thing."

Another useful way of thinking about energy is the following:

Energy is what is required in order to resist a natural tendency.

For example, consider holding a bowling ball above your head. You are not moving the ball, so there is no work. But obviously you get tired after awhile, so it feels as though you are exerting energy. How can this be? The natural tendency of the bowling ball is to fall to the ground (due to gravity). By keeping the ball from falling, you are making the ball resist its natural tendency; therefore you *are* exerting energy. The same goes if you are holding a string connected to a helium balloon. The natural tendency of the balloon is to float away, and by holding the string you are keeping the balloon from doing so. This, too, requires energy.

In discussing energy, you also need to keep distinct potential energy and kinetic energy. The "ball on a hill" example given in the text is a good one to read and understand, but it is a non-chemical example. When thinking about chemistry, keep in mind that the potential energy of a chemical system is stored in the bonds (you will learn more about chemical bonds in Chapter 12, but it is enough for now to know that a chemical bond is a force that holds atoms together in a molecule). In order to "break apart" a molecule, energy is required to break the bonds. As bonds reform when new molecules are made, energy is released in the form of heat (as kinetic energy).

Let's look at the reaction that occurs in a Bunsen burner (the combustion of methane):

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

In this case, more energy is released when the bonds form CO_2 and H_2O than is required to break the bonds in CH_4 and O_2. Thus this reaction is exothermic, and we report the heat with a negative sign (in this case ($\Delta H = -891$ kJ/mol). See Figure 10.5 in your text for a potential energy diagram of this reaction. It would be a good idea to see if you can draw a similar diagram for an endothermic reaction.

You should also be able to differentiate between heat and temperature. Specifically, make sure you understand that heat and temperature are not the same. Temperature is a measure of the random motions of the particles that make up a substance. The concept of specific heat capacity helps us see this difference. You undoubtedly have seen that different substances change temperatures differently when the same amount of heat is transferred. For example, if you are making soup on the stove and are stirring with a metal spoon, you notice that the spoon gets hot rather quickly. A wooden spoon does not get nearly so hot, even though the amount of heat is the same. Why is this? Different substances, due to their make-ups, react to heat differently. We quantify this with the specific heat capacity, which is defined as the amount of energy required to change the temperature of one gram of the substance by one Celsius degree. See Table 10.1 in your text for a list of specific heat capacities, and notice that the heat capacities of metals are lower than water. This means that the temperature of a given mass of metal will increase much more than the temperature of the same mass of water when the same amount of heat is transferred to each. Practice doing problems dealing with specific heat capacities to make sure you can either use heat capacities to calculate temperature differences or that you can calculate heat capacities and determine the substance that is being heated or cooled.

You will also be expected to calculate the heat (enthalpy) of a reaction from known heats of related reactions. You can do this using Hess's law, which works because energy is a state function. Make sure you understand how Hess's law uses the idea that energy is a state function (another way of stating this is to rely on the existence of the first law of thermodynamics – make sure you know why).

While energy is conserved (that is the quantity stays the same), the quality of energy is constantly decreasing. That is, the amount of usefulness in a sample of energy decreases as the energy is "used." The example used in your text is gasoline. A sample of gasoline can be thought of as concentrated energy (potential energy). However, as you drive the car heat is released to the road, the air, etc., and so, while the amount of energy in the universe is the same before and after you drive your car, the usefulness has decreased. Natural processes always occur in a way that increases the "spreading" of energy and thus decreases the usefulness of the energy. We term this entropy, which is a measure of disorder. For a process to be spontaneous, the entropy of the universe must increase.

LEARNING REVIEW

1. Explain the difference between *kinetic energy* and *potential energy*.
2. Why isn't all energy available as work?
3. The law of conservation of energy means that energy is a state function. Explain why.
4. Explain differences among heat, temperature, and thermal energy.
5. Provide a molecular-level explanation of why the temperatures of a cold soft drink and hot coffee in the same room will eventually be the same.
6. In which case is more heat involved: mixing 100.0-g samples of 90 °C water and 80 °C water or mixing 100.0-g samples of 60 °C water and 10 °C water? Assume no heat is lost to the environment.
7. What is meant by potential energy in a chemical reaction? Where is it located?
8. Are the following processes exothermic or endothermic?
   a. When solid KBr is dissolved in water, the solution gets colder.
   b. Natural gas (CH_4) is burned in a furnace.
   c. When concentrated sulfuric acid is added to water, the solution gets very hot.
   d. Water is boiled in a tea kettle.

9. In thermodynamics the chemist takes the system's point of view. What does this statement mean?

10. A gas absorbs 45 kJ of heat and does 29 kJ of work. Calculate ΔE.

11. Convert the energy values below to the desired units.

    a. 45.8 cal to J

    b. 0.561 cal to J

    c. 5.96 J to cal

    d. 76 J to cal

12. Calculate the number of calories required to change the temperature of each of the quantities of water below.

    a. 100.1 g of water from 6°C to 25°C

    b. 2.32 g of water from 36°F to 42°F

    c. 40 g of water by 12°C

    d. 16.9 g of water from 75.0°C to 80.0°C

13. How much energy (in joules) is required to raise the temperature of 25.2 g of solid carbon rod from 25 °C to 50.°C? The specific heat capacity of solid carbon is 0.71 J/g °C.

14. How much energy (in calories) is required to raise the temperature of 10. g steam from 122.2 °C to 130.4 °C? The specific heat capacity of water(*g*) is 2.0 J/g °C.

15. How much of a temperature change would occur if 2736.8 J of energy were applied to a piece of iron bar weighing 450.5 g? The specific heat capacity of solid iron is 0.45 J/g °C.

16. What is the mass in grams of a piece of aluminum wire if a change in temperature of 5.67 °C required 8.53 J? The specific heat capacity of solid aluminum is 0.89 J/g °C.

17. What is the specific heat capacity of ethyl alcohol if 1972.4 J of energy is necessary to raise the temperature of 53.4 g ethyl alcohol by 15.2°C?

18. Calculate the enthalpy change when 1.00 g of methane is burned in excess oxygen according to the reaction $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + H_2O(l)$ ($\Delta H = -891$ kJ/mol).

19. Given the following data:

$$C_2H_2(g) + \tfrac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(l) \qquad \Delta H = -1300.\ \text{kJ/mol}$$
$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -394\ \text{kJ/mol}$$
$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l) \qquad \Delta H = -286\ \text{kJ/mol}$$

calculate ΔH for the reaction

$$2C(s) + H_2(g) \rightarrow C_2H_2(g)$$

20. What is the difference between the *quality* of energy and the *quantity* of energy? Which is decreasing?

21. Which energy sources used in the United States have declined the most in the last 150 years? Which have increased the most?

22. Why can't the first law of thermodynamics explain why a ball doesn't spontaneously roll up a hill?

23. Exothermic reactions have a driving force. Nevertheless, water melting into a liquid is endothermic, and this process occurs at room temperature. Explain why.

ANSWERS TO LEARNING REVIEW

1. Kinetic energy is the energy of motion. Potential energy is the energy of position.
2. Some energy is given off in other forms such as heat or light.
3. Because energy is conserved in any process, the pathway of the process does not matter. This means that energy is a state function.
4. Heat is a flow of energy due to a temperature difference; temperature is a measure of the average kinetic energy of a substance; thermal energy comes from the random motion of the components of the system.
5. Temperature is a measure of the average kinetic energy of the samples. The "coffee particles" (mostly water molecules) are of higher average kinetic energy than the "air particles" (a mixture of mostly nitrogen and oxygen molecules) in the room, which are of higher energy than the "soft drink particles" (mostly water molecules). At the coffee-air interface, a collision of higher-energy "coffee particle" and "air particle" results in energy being transferred from the coffee to the room. Transfer also occurs from air to soft drink because "air particles" are of higher energy than the "soft drink particles" (due to the temperature difference). The energy transfers result in the eventual average kinetic energies of each sample being equal, which means the temperatures are equal. Because the volume of air is so large (the system is open), however, no noticeable temperature change of the air will result.
6. There is more heat involved in mixing 100.0-g samples of 60 °C and 10 °C water because there is a large temperature difference (and heat is a flow of energy due to a temperature difference).
7. The potential energy is the energy available to do work. Potential energy in a chemical reaction is stored in the chemical bonds.
8.
   a. endothermic
   a. exothermic
   b. exothermic
   c. endothermic
9. The chemist chooses the sign based on whether energy flows from the system (negative sign) or into the system (positive sign).
10. $\Delta E = q + w = 45\text{ kJ} + (-29\text{ kJ}) = 16\text{ kJ}$
11. Converting calories to joules or joules to calories requires knowing that 1 cal = 4.184 J.

   a. $45.8\text{ cal} \times \dfrac{4.184\text{ J}}{1\text{ cal}} = 192\text{ J}$

   b. $0.561\text{ cal} \times \dfrac{4.184\text{ J}}{1\text{ cal}} = 2.35\text{ J}$

   c. $5.96\text{ J} \times \dfrac{1\text{ cal}}{4.184\text{ J}} = 1.42\text{ cal}$

d. $76\ \text{J} \times \dfrac{1\ \text{cal}}{4.184\ \text{J}} = 18\ \text{cal}$

12. These problems ask you to calculate the calories required to heat a quantity of water. One calorie is defined as the amount of heat required to raise the temperature of 1 gram of water by 1 degree Celsius. To solve these problems, you need to multiply the number of grams of water to be heated by the number of degrees Celsius change in the temperature of the water.

a. $100.1\ \text{g water} \times 19°\text{C} \times \dfrac{1\ \text{cal}}{\text{g water} \times °\text{C}} = 1900\ \text{cal}$

b. The initial and final temperatures of water are given in °F. We must convert to °C before solving the problem.

Initial temperature:

$$T_{°C} = \frac{(T_{°F} - 32)}{1.80}$$

$$T_{°C} = \frac{(36 - 32)}{1.80}$$

$$T_{°C} = 2.2$$

Final temperature:

$$T_{°C} = \frac{(T_{°F} - 32)}{1.80}$$

$$T_{°C} = \frac{(42 - 32)}{1.80}$$

$$T_{°C} = 5.6$$

Temperature change:

$$5.6\ °\text{C} - 2.2\ °\text{C} = 3.4\ °\text{C}$$

Solution: $2.32\ \text{g water} \times 3.3°\text{C} \times \dfrac{1\ \text{cal}}{\text{g water} \times °\text{C}} = 7.9\ \text{cal}$

c. $40.\ \text{g water} \times 12°\text{C} \times \dfrac{1\ \text{cal}}{\text{g water} \times °\text{C}} = 480\ \text{cal}$

d. $16.9\ \text{g water} \times 5.0°\text{C} \times \dfrac{1\ \text{cal}}{\text{g water} \times °\text{C}} = 85\ \text{cal}$

13. In this problem we want to calculate the heat energy needed to raise the temperature of a substance other than water. To do this, we need to know the specific heat capacity of the substance. The specific heat capacity tells us the amount of heat energy required to change the temperature of 1 gram of a substance by 1 degree Celsius. Every substance has its own specific heat capacity. That of solid carbon is 0.71 J/g °C.

If it takes 0.71 J to raise the temperature of 1 gram of carbon 1 degree Celsius, then it will take 0.71 J × 25.2 to raise 25.2 g carbon by 1 degree Celsius. We wish to raise the temperature of the carbon rod by 25 °C, not 1 °C. We will need twenty-five times the heat energy needed to raise the temperature of 25.2 g carbon by 1 degree Celsius.

$$\text{Joules} = \frac{0.71\text{ J}}{\text{g carbon} \times {}^\circ\text{C}} \times 25.2\text{ g carbon} \times 25^\circ\text{C}$$

The joules required to raise the temperature of 25.2 g carbon by 25 °C = 450 J.

14. When you are given the number of grams of a substance, a change in temperature in degrees Celsius, and the specific heat capacity for that substance and are asked to calculate the heat energy required, you can use the formula $Q = s \times m \times \Delta T$. The specific heat capacity is s, m equals the mass in grams, ΔT is the change in temperature in degrees Celsius, and Q is the heat energy required. We can solve this problem with the formula although we will need to convert Q from joules to calories since calories are asked for.

$$Q = s \times m \times \Delta T$$

$$Q = \frac{2.0\text{ J}}{\text{g water}(g) \times {}^\circ\text{C}} \times 10.\text{ g water}(g) \times 8.2\ {}^\circ\text{C}$$

$$Q = 160\text{ J}$$

The answer should be expressed in calories:

$$160\text{ J} \times \frac{1\text{ cal}}{4.184\text{ J}} = 38\text{ cal}$$

15. In this problem, we are given Q, the heat energy in joules; m, the mass in grams of a piece of iron; and s, the specific heat capacity of iron. We are asked for ΔT, the change in temperature. If we rearrange the equation $Q = s \times m \times \Delta T$, we can solve for ΔT.

Divide both sides by ΔT.

$$\frac{Q}{\Delta T} = s \times m$$

Now, divide both sides by Q (same as multiplying by $\frac{1}{Q}$).

$$\frac{Q}{\Delta T} \times \frac{1}{Q} = \frac{s \times m}{Q}$$

We now have $1/\Delta T$ (the inverse of ΔT) isolated on one side of the equation.

$$\frac{1}{\Delta T} = \frac{s \times m}{Q}$$

Invert both sides of the equation.

$$\frac{\Delta T}{1} = \frac{Q}{s \times m}$$

$$\frac{\Delta T}{1} = \Delta T = \frac{Q}{s \times m}$$

Now, find ΔT.

$$\Delta T = \frac{2736.8 \text{ J}}{\frac{0.45 \text{ J}}{\text{g °C}} \times 450.5 \text{ g}}$$

$$\Delta T = 14 \text{ °C}$$

16. In this problem we are asked to solve for mass, m. We are given ΔT, Q, and s. We can rearrange the equation as illustrated below.

Divide both sides of the equation by m.

$$\frac{Q}{m} = \frac{s \times m \times \Delta T}{m}$$

Divide both sides of the equation by Q.

$$\frac{Q}{m} \times \frac{1}{Q} = s \times \Delta T \times \frac{1}{Q}$$

$$\frac{1}{m} = \frac{s \times \Delta T}{Q}$$

Invert both sides of the equation.

$$\frac{m}{1} = m = \frac{Q}{s \times \Delta T}$$

Now, find m.

$$m = \frac{8.53 \text{ J}}{\frac{0.89 \text{ J}}{\text{g} \times \text{°C}} \times 5.67 \text{ °C}}$$

$$m = 1.7 \text{ g}$$

17. We are asked to find the specific heat capacity, s, when given ΔT, Q, and m. Rearrange the equation to isolate s on one side of the equation.

$$s = \frac{Q}{m \times \Delta T}$$

Now substitute values into the equation.

$$s = \frac{1972.4 \text{ J}}{54 \text{ g} \times 15.2 \text{ °C}} = 2.43 \text{ J/g °C}.$$

The specific heat capacity of ethyl alcohol is 2.43 J/g °C.