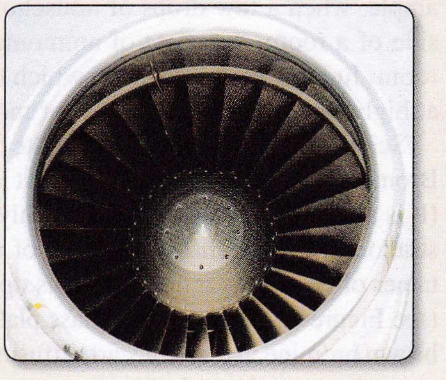
Thermal Energy and Heat

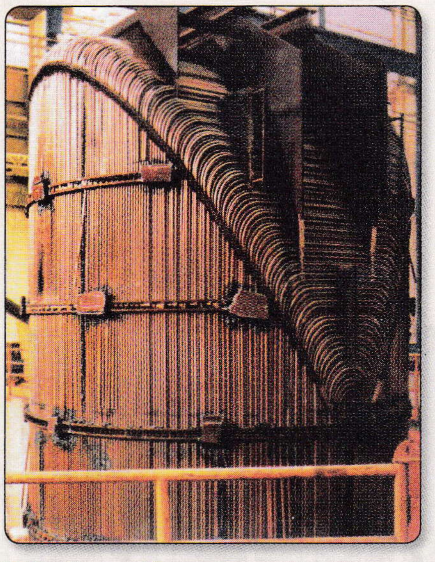
15

Dominion Modeling

Metals and Heat

Why are different types of metals used for different purposes? Specific metals and mixtures of metals, called alloys, are used for cookware, jet engines, steam boilers, catalytic converters, and many other applications. Metals are good conductors of heat and electricity and generally possess strength, ductility, and malleability. But why do different metals have different thermal properties? Can this difference be related to other measurable properties of metals?





15-1 Why are different metals used for different purposes?

15A THEORIES OF HEAT

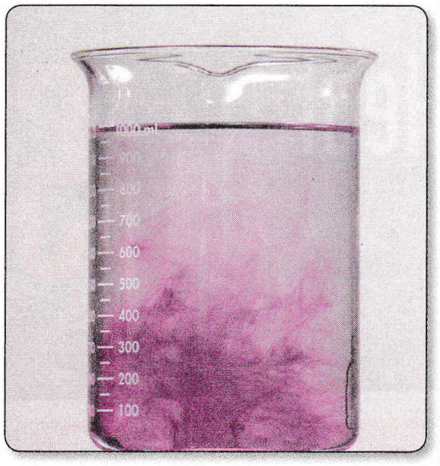
15.1 Introduction

What makes one object hot and another cold? What exactly is heat? Scientists have studied these questions for centuries, and their conclusions have changed as they have learned more about how heated and unheated objects behave.

15.2 The Kinetic Theory

Modern scientists agree that what we call *heat* is really the movement of a form of energy called thermal energy. According to the kinetic theory of matter, all substances contain tiny, constantly moving particles. The particles in hot materials move faster than the particles in colder materials. **Thermal energy** is the sum of the kinetic energy of the random motion of the particles. The average kinetic energy of the particles is proportional to the temperature of the material. However, temperature alone does not indicate how much thermal energy is present. We will look into this aspect later. For now, let's review some important observations that give evidence for the kinetic-molecular theory of matter.

The kinetic theory is also called the kinetic-molecular theory of matter because the two concepts of tiny discrete particles in constant motion are inseparable.



15-2 The spreading of this dye throughout the water illustrates the principle of diffusion.

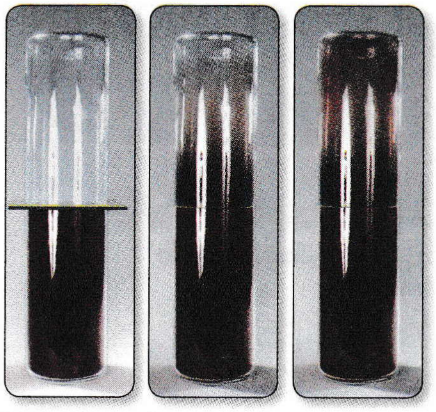
15.3 Matter Can Be Subdivided

Suppose that you place a crystal of blue dye in a beaker of water and stir it until it is completely dissolved and all the water has become dark blue. The crystal has somehow divided itself and spread throughout the water in the beaker. If you pour the contents of the beaker into a much larger container of water and stir the solution again, the color becomes less intense but is still completely distributed throughout the larger container. The process may be repeated many times. The color becomes progressively fainter but is still present. The material from the crystal seems to keep dividing indefinitely. For this to be true, there must be a multitude of minuscule particles. These particles are called molecules, atoms, or ions.

15.4 Diffusion

You have seen evidence that particles exist. Now let's consider evidence that these particles are in motion. Suppose you drop a crystal of the same blue dye into a beaker of water *without* stirring. If you wait patiently, you will see that the crystal dissolves and that eventually the dye and the water completely mix (though this may take a very long time). This observation can be explained by assuming that the water molecules are in constant motion. Their motion distributes particles of the dye throughout the water. The spreading of a substance through particle motion alone is called **diffusion**.

Diffusion takes place much faster in gases than in liquids. This observation leads us to believe that gas molecules move faster than liquid molecules. For example, when a few drops of household ammonia are spilled from a bottle on one side of a room, the scent of ammonia can soon be smelled on the far side of the room. Judging by the speed at which the ammonia drifts, air molecules must move at high speeds, around 500 m/s or more at room temperature.



15-3 Diffusion of gases is clearly seen in the mixing of bromine vapor and air.

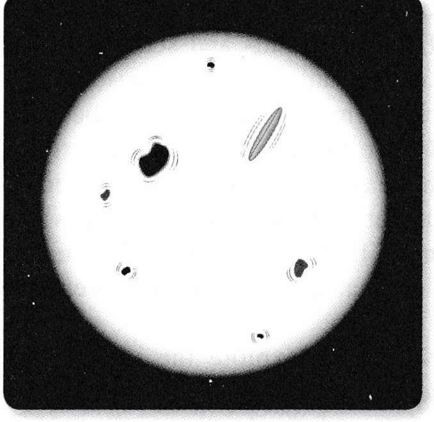
Scientists can visually demonstrate diffusion in gases using bromine vapor and air. Bromine vapor is brown and dense; it is about five and one-half times as dense as air. (It is also poisonous—that is why only qualified people should perform this demonstration.) Suppose that a container of air is placed upside down over a similar container of bromine vapor that has its open end facing up but covered by a glass plate (see Figure 15-3). When the glass plate is slid out from between the containers, the brown bromine vapor will start diffusing upward into the transparent air. Air will also move downward into the bromine vapor. This cannot be explained by gravity, because the lighter air goes downward, while the heavier bromine vapor moves upward. After a period of time, the two gases will be evenly mixed. The molecules of the air and the bromine must each have motion of their own to mix themselves so quickly and thoroughly. This process occurs even more rapidly at higher temperatures.

Robert Brown (1773–1858) was an English botanist who first described Brownian motion while studying pollen grains under a microscope.

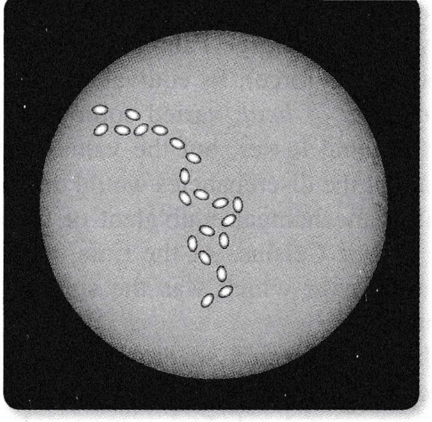
15.5 Brownian Motion

In 1827 Robert Brown discovered more direct evidence that matter is in constant motion. Brown used a microscope to examine plant spores floating in water. He saw the spores jostling back and forth as if they were being struck repeatedly from different sides. Yet nothing was touching them except the water in which they were floating. The explanation for his observation is that water molecules are in constant random motion, and they collide with the spores. This effect, seen only in tiny particles, is called **Brownian motion**.

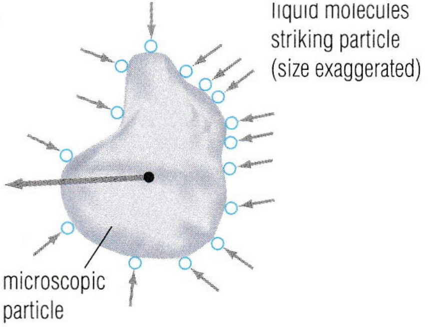
It is also possible to see Brownian motion in smoke particles. While looking through a microscope, introduce a small amount of smoke before the objective lens and illuminate it with a bright light. You will be able to clearly see the smoke particles jostling about in vibratory motion. This motion is caused by collisions with high-speed air molecules.



15-4a Brownian motion of microscopic particles suspended in a liquid or gas can be observed using a light microscope.



15-4b Time-lapse microphotography shows the position of an inanimate particle experiencing Brownian motion at given time intervals.



15-4c Brownian motion of microscopic particles is caused by the random, asymmetrical collisions of liquid or gas molecules against the particles.

15.6 The Caloric Theory

Many eighteenth-century scientists believed the **caloric theory**—the view that heat is a material fluid, called *caloric*, that flows from hot bodies to cold bodies. Earlier scientists, including Francis Bacon, Robert Boyle, and Robert Hooke, had previously advocated the rival kinetic theory—the theory that thermal energy is a property of an object arising from the motion of its particles. Still, such notable scientists as Joseph Black, who discovered carbon dioxide, and Antoine Lavoisier, the famous French chemist, enthusiastically endorsed the caloric theory.

Joseph Black (1728–99) was a Scottish physician who conducted research in the thermal properties of matter and discovered carbon dioxide. Black also proposed the concepts of latent and specific heat. His work was foundational in the establishment of the science of chemistry.

Antoine Lavoisier (1743–94) was a brilliant French chemist often called the "Father of Modern Chemistry." He discovered the conservation of mass and identified oxygen as the active agent in combustion and respiration. He was beheaded in the French Revolution.

A major blow to the caloric theory came with the famous cannon-boring experiment conducted by Count Rumford. According to the caloric theory, the amount of thermal energy produced in drilling a metal should be proportional to the volume of the material the drill removed. The caloric fluid was believed to reside between the most elementary particles of the metal and be released when the particles were separated.

However, Rumford observed no correlation between the volume of the material removed and the thermal energy produced. On the contrary, he observed that the greatest heat was produced when a dull drill removed only a little metal after a great amount of mechanical work was expended. An almost unlimited supply of thermal energy could be produced under such conditions. In a report to the Royal Society in 1798, he theorized that thermal energy was the result of motion in the elementary particles of the metal. Humphry Davy continued this line of attack on the caloric theory in the early 1800s.

Humphry Davy (1778–1829) was an English chemist who specialized in discovering many new elements and investigating their properties. He discovered the law of combining proportions that supported Dalton's atomic theory of matter. He also conducted qualitative research in many areas of science.

In 1840 James Prescott Joule of Manchester, England, began a series of experiments designed to unite thermal and mechanical energy under a single conservation law. In a lecture that he delivered in 1847, Joule administered the *coup de grâce* to the faltering caloric theory. He showed that since caloric can be converted to kinetic or potential energy, it cannot possibly be a material substance. He explained the source of the heating in Rumford's experiment as a change from mechanical to thermal energy. The caloric theory quickly became a thing of the past. The paradigm shift from the caloric to the kinetic-molecular model of thermal energy and heat is one of the most profound scientific revolutions in history because it allowed the science of thermodynamics to develop rapidly.

15.7 The Mechanical Equivalent of Heat

The results of Joule's research, published in 1845, described a device in which falling weights turned a set of paddles in an enclosed drum of water. After conducting a series of careful experiments, he reported that it required an average of 819 foot-pounds-force (ft-lbf) of work to raise the temperature of a pound of water one degree Fahrenheit (he later refined his results and settled on 772 foot-pounds-force). Its equivalent in thermal energy units, called the *mechanical equivalent of heat*, varied somewhat from experiment to experiment because of frictional losses, but the values obtained were close enough to support the idea that the discrepancies would disappear in the absence of friction. In SI units, Joule's mechanical equivalent of heat was about 4.50 N·m per gram of water per degree Celsius. At the time, this amount of heat was assigned the unit of calorie (cal), which was the standard unit of heat until late in the twentieth century.

The calorie (cal) was the unit of heat for many years. However, the joule is now the primary unit for both energy and heat in most scientific research. The calorie, or more specifically, the dietary Calorie (1 Calorie = 1 kcal or 1000 cal), is still in use for quantifying the energy content of foods.

Joule performed several different experiments to explore the idea of the mechanical equivalent of heat. His experiments did not agree exactly. For the equivalent of 1 cal, Joule reported 4.14 N·m of work for heating water with a perforated piston, 4.27 N·m for the temperature drop of rapidly expanding air, and 4.16 N·m for experiments involving water moving through narrow tubes. By 1850 he had settled on a value of 4.15 N·m for 1 cal. The currently accepted value of the mechanical equivalent of thermal energy is 4.186 N·m = 1 cal (at 15 °C). In honor of the fundamental work Joule accomplished in thermodynamics, the N·m was renamed the joule, the SI derived unit of energy, work, and heat.

15A Objectives

After completing this section, I can

- ✓ summarize the evidence for the kinetic-molecular model of matter.
- ✓ distinguish between diffusion and Brownian motion.
- ✓ discuss the basis of the caloric theory and its demise in the face of Joule's discovery of the mechanical equivalent of heat.
- ✓ compare the different units of heat and energy.

15A Section Review

1. According to the kinetic theory, describe how thermal energy is manifested in matter.
2. How is temperature related to thermal energy?
3. Discuss three evidences that support the kinetic-molecular theory of matter.
4. What observation produced a key piece of evidence that contradicted the caloric theory of heat and led to its eventual demise?
5. What logical argument proved that heat could not be a material fluid called *caloric*?
6. State the SI equivalent for 1 cal.
7. Ignoring losses to the surroundings, if 10.0 J of work is done stirring 2.0 L of water, how much thermal energy is transferred to the water?

15B THERMAL ENERGY AND MATTER

15.8 Heat Capacity

Recall from earlier chapters that momentum and velocity are related but different quantities. It is not always the fastest-moving object that has the greatest momentum. For example, a slow-moving truck may have more momentum than a speeding motorcycle because the truck has far more mass than the motorcycle. Momentum is the product of mass and velocity.

The heat capacity (C) of an object is the quantity of thermal energy required to raise the temperature of the *entire object* 1 degree Celsius. Units are J/°C. Heat capacity is a unique property of each object.

Heat (Q) is the quantity of thermal energy that is transferred to or from a system. It is not correct to say that an object contains a certain amount of heat.

Similarly, it is not always the hottest object that contains the greatest amount of thermal energy. There are two other factors involved here—the mass of the object and its ability to hold thermal energy. Both of these factors are included in the object's heat capacity (C), which is the amount of thermal energy required to raise the temperature of the entire object one degree Celsius. The amount of thermal energy added to or taken from a system is called heat (Q). The SI unit for heat capacity is joules per degree Celsius (J/°C). Heat capacity of an object is to thermal energy as mass is to momentum. Heat capacity may be calculated by the formula

$$C = \frac{Q_{\text{object}}}{\Delta t}, \qquad (15.1)$$

where Δt is the change of temperature of the object,

$$\Delta t = t_{\text{final}} - t_{\text{initial}}.$$

Heat is analogous to mechanical work. Mechanical work on or by a system involves the transfer of mechanical energy to or from the system. Heat involves the movement of thermal energy. So, it is technically incorrect to say that a system *has* a certain amount of heat, just as it is improper to say that a system *has* a certain amount of work. Heat is not a property of a system as thermal energy is. It takes great discipline to use the two terms correctly.

Just as momentum is a function of a fixed property of an object (its mass) and a variable property (its speed), the thermal energy of an object is dependent on its heat capacity (a fixed property) and its temperature (a variable property).

15.9 Specific Heat

Mass is a property of a specific object, not of a kind of substance. Similarly, heat capacity is a property of an object. To compare the relative "heaviness" of one substance to another, scientists divide the mass of a sample of the substance by its volume, which yields the *specific density* of the material—its mass per unit volume. Similarly, to find the thermal energy capacity of a substance (not an object), scientists divide the heat capacity of an object by its mass to find its *specific heat capacity*—its heat capacity per gram. The **specific heat (c_{sp})** of a substance is the amount of thermal energy required to raise the temperature of 1 g of the substance one degree Celsius. It is properly expressed in SI units of joules per gram-degree Celsius (J/g·°C). Older textbooks sometimes refer to units of calories per gram-degree Celsius. Comparing the definitions of specific heat and calorie, we can see that the specific heat of water is, by definition, exactly 1 cal/g·°C (at 15 °C). In SI units, the specific heat of water is about 4.18 J/g·°C (near room temperature). The specific heats of some common substances are listed in Table 15-1 on page 344. Specific heat varies somewhat with temperature and pressure. However, the specific heats of most solids are nearly constant for temperatures between 20 °C and 100 °C and pressures near 1 atm.

Recall that lowercase *t* refers to Celsius temperature as well as time. This is another case in physics where a single symbol is used for more than one quantity. The context of the problem will indicate what dimensional property is meant.

Specific heat capacity, or just **specific heat (c_{sp})**, is the amount of heat required to raise the temperature of 1 g of a substance 1 °C.

The definition of the calorie, dependent as it is on the heat capacity of water (which varies with temperature), must be given with an associated temperature. Typically, the calorie is defined at 15 °C, although other temperatures are used as well.

How can we find the specific heat of a material? Using the definition of specific heat, if we add a known amount of thermal energy (Q) to an object of known mass (m) and note the corresponding temperature change (Δt) in degrees Celsius, then we can find the specific heat of the material from the equation

$$c_{sp} = \frac{Q}{m\Delta t}. \qquad (15.2)$$

EXAMPLE 15-1

Taking Some Heat: Determining Heat Capacity

When 89.7 J of thermal energy is added to a 10.0 g aluminum block at 20.0 °C, the block's temperature rises to 30.0 °C. Compute (a) the specific heat of aluminum and (b) the heat capacity of the block.

Solution:

The system is the aluminum block.

a. Computing specific heat:

$$c_{sp\ Al} = \frac{Q}{m\Delta t} = \frac{89.7\ \text{J}}{(10.0\ \text{g})(10.0\ ^\circ\text{C})}$$

$$c_{sp\ Al} = 0.897\ \text{J/g}\cdot{}^\circ\text{C}$$

Equation 15.1 can be used to find the heat capacity of both objects and substances. However, to find the heat capacity of uniform substances, it is better to multiply the substance's specific heat by its mass. Example 15-1 takes this approach in Part b.

b. Computing heat capacity:

$$C = mc_{\text{sp Al}}$$

$$C = (10.0\ \text{g})(0.897\ \text{J/g}\cdot{}^{\circ}\text{C})$$

$$C = 8.97\ \text{J/}^{\circ}\text{C}$$

The amount of heat that is needed to raise the temperature of an object a certain number of degrees is released when the object is cooled the same number of degrees. If, in the previous example, the aluminum block were cooled from 30 °C to 20 °C, the cooling process would release 89.7 J of heat.

15.10 Modeling Specific Heat

Compare the specific heats of silver and gold in Table 15-1. Both are metals and both have free electrons, which cause them to be good conductors of heat. Both metals naturally exist in the same kind of crystal lattice structure. When these metals are heated, atoms in the lattice can only vibrate in place, since metal bonds keep atoms from significantly changing position. So, if the two metals' particle arrangements and thermal motions are nearly identical, what accounts for the difference in the value of specific heat of these two metals? It must have to do with some difference in the metal atoms themselves. An element's measurable properties are mainly due to the mass and electron structure of its atoms. Since the electron structures of the two metals are similar, it seems reasonable that there must be some relationship between an element's mass (atomic weight) and its specific heat value. Look at the following table that gives the values of the specific heat and **molar masses** of several pure metals. Later, you will be asked to make a model relating these two quantities.

TABLE 15-2

Physical Data for Selected Metals

Metal	Molar Mass (g/mol)	Specific Heat (J/g·°C) at 25 °C
Mg	24.31	1.024
Al	26.98	0.897
Cr	52.00	0.450
Fe	55.85	0.449
Co	58.93	0.421
Cu	63.55	0.384
Zn	65.41	0.388
Ag	107.86	0.235
W	183.84	0.132
Au	196.97	0.129
Pb	207.21	0.127

Ref: *Handbook of Chemistry and Physics*, 86th ed. David R. Lide, Ed. CRC Press. 2005

TABLE 15-1

Specific Heat of Selected Substances

Substance	c_{sp} (J/g·°C)
air (27 °C/300 K)	1.16
aluminum	0.897
benzene	1.75
brass	0.375
carbon (graphite)	0.709
chlorine	0.479
copper	0.385
ethyl alcohol	2.44
glass (crown)	0.670
gold	0.129
iron	0.449
lead	0.129
mercury	0.140
oil (olive)	1.79
oxygen	0.918
platinum	0.133
silver	0.235
tin	0.228
tungsten	0.132
water (solid/ice)	2.09
water (liquid)	4.18
water (gas/steam)	2.01
wood	1.70
zinc	0.388

The free electrons found in metals are unlocalized valence electrons involved in forming the bonds between metal atoms. We will discuss this further in Subsection 15.13.

The particles in liquids and gases not only oscillate, but they also can rotate and move along linear paths (translation). Molecules also exhibit complicated internal motions among their atoms. Thus, the kinetic energies associated with fluid particles tend to have a complex effect on their measurable properties.

15.11 Conservation of Thermal Energy

When an object gains heat, its surroundings lose the same amount of heat. Thermal energy is conserved if it is the only form of energy exchanged in a process. In this chapter we will assume that thermal energy is conserved, so the heat gained by a system is lost by its surroundings and vice versa. Symbolically,

$$Q_{\text{system}} = -Q_{\text{surroundings}}, \text{ or}$$
$$Q_{\text{system}} + Q_{\text{surroundings}} = 0 \text{ J.}$$

Such expressions are called *heat-balance equations.*

Problem-Solving Strategy 15.1

Heat-balance problems often result in slightly different calculated values of heat lost versus heat gained. This is due to the different precisions of the various measured quantities and the rounding conventions involved in the calculations. Differences of a few percent are acceptable for the purposes of this course.

An **adiabatic** process is one that does not exchange thermal energy with its surroundings.

EXAMPLE 15-2

Conservation of Thermal Energy: The Calorimeter

A 10.0 g aluminum block has an initial temperature of 19.7 °C. It is placed in an **adiabatic** vessel—one that allows no heat to enter or leave its contents—that contains 20.0 g of water at 31.1 °C. The two parts of the system come to equilibrium at a final temperature of 30.0 °C. Compare the heat gained by the metal block to the heat lost by the water.

Solution:

The system consists of the water and the aluminum block. It does *not* include the vessel for this problem. You are seeking the quantity of heat lost and gained by the respective parts of the system. Equation 15.2 must be rearranged to solve for Q:

$$Q = mc_{sp}\Delta t$$

Calculate the heat gained by the water:

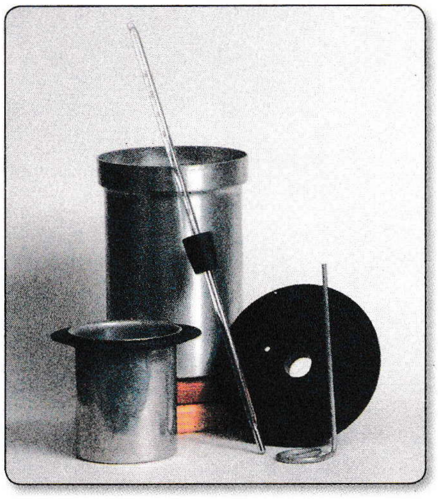
$$Q_{H_2O} = (mc_{sp}\Delta t)_{H_2O} = [mc_{sp}(t_f - t_i)]_{H_2O}$$
$$Q_{H_2O} = (20.0 \text{ g})(4.18 \text{ J/g}\cdot°\text{C})(30.0 \text{ °C} - 31.1 \text{ °C})$$
$$Q_{H_2O} \cong -9\underline{1}.9 \text{ J} \cong -92 \text{ J} \qquad (1)$$

The heat "gained" by the water is negative, so thermal energy is actually lost from the water. There is a net flow of thermal energy out of the water and into the aluminum block.

Calculate the heat gained by the block:

$$Q_{Al} = (mc_{sp}\Delta t)_{Al} = [mc_{sp}(t_f - t_i)]_{Al}$$
$$Q_{Al} = (10.0 \text{ g})(0.897 \text{ J/g}\cdot°\text{C})(30.0 \text{ °C} - 19.7 \text{ °C})$$
$$Q_{Al} \cong +92.\underline{3}9 \text{ J} \cong +92.4 \text{ J} \qquad (2)$$

Notice that $Q_{Al} = -Q_{H_2O}$ or $+92 \text{ J} = -(-92 \text{ J})$ expressed to 2 SDs.



15-5a Components of a calorimeter

15-5b A research-grade calorimeter

In the previous example, an adiabatic container was specified. There are no truly adiabatic containers. Therefore, we must take into account the thermal energy lost or gained by the container and the surrounding air. A **calorimeter** is a container designed to minimize the exchange of thermal energy with its surroundings. It consists of an insulated outer cup, an inner cup, an insulated lid, and provision for a stirrer and a thermometer. Figure 15-5a is a photograph of a typical school laboratory calorimeter. Figure 15-5b is a research-quality calorimeter. It is possible to determine the heat capacity of a calorimeter by using the principle of the conservation of energy.

The title of Example 15-3 refers to Song of Solomon 2:15: "Take us the foxes, the little foxes, that spoil the vines"—the little foxes are an allegorical reference to the small iniquities in our lives that rob us of our joy in our Lord just as the calorimeter affects, if only in a small way, the thermal energy transferred between its contents.

EXAMPLE 15-3

Finding the Little Foxes: Calorimeter Heat Capacity

10.0 g of water at room temperature is placed into a calorimeter. The equilibrium temperature of the calorimeter and water is 20.0 °C. Then 10.0 g of water heated to 40.0 °C is added to the water in the calorimeter. The new equilibrium temperature is 27.0 °C. Find the heat capacity of the calorimeter.

Solution:

The system is the calorimeter, the cool water, and the additional hot water. You measure the water temperature and assume that the calorimeter's temperature is the same at equilibrium. The essential concept at work here is this: After the hot and cool water mix and you determine the heat gained and lost by the two quantities of water, any difference in heat *must* be due to the heat transferred to or from the calorimeter, based on the conservation of thermal energy.

$$Q_{\text{hw}} + Q_{\text{cw}} + Q_{\text{cal}} = 0\ \text{J} \tag{1}$$

Calculate the heat transfer of the cool water:

$$Q_{\text{cw}} = (mc_{\text{sp}}\Delta t)_{\text{cw}} = [mc_{\text{sp}}(t_{\text{f}} - t_{\text{i}})]_{\text{cw}}$$

$$Q_{\text{cw}} = (10.0\ \text{g})(4.18\ \text{J/g}\cdot{}^{\circ}\text{C})(27.0\ {}^{\circ}\text{C} - 20.0\ {}^{\circ}\text{C})$$

$$Q_{\text{cw}} \cong +2\underline{9}2\ \text{J} \tag{2}$$

Calculate the heat transfer of the hot water:

$$Q_{\text{hw}} = (mc_{\text{sp}}\Delta t)_{\text{hw}} = [mc_{\text{sp}}(t_{\text{f}} - t_{\text{i}})]_{\text{hw}}$$

$$Q_{\text{hw}} = (10.0\ \text{g})(4.18\ \text{J/g}\cdot{}^{\circ}\text{C})(27.0\ {}^{\circ}\text{C} - 40.0\ {}^{\circ}\text{C})$$

$$Q_{\text{hw}} \cong -54\underline{3}.4\ \text{J} \tag{3}$$

Substitute results (2) and (3) into Equation (1) and solve for the heat transfer of the calorimeter:

$$Q_{\text{hw}} + Q_{\text{cw}} + Q_{\text{cal}} = 0\ \text{J}$$

$$Q_{\text{cal}} = -(Q_{\text{hw}} + Q_{\text{cw}})$$

$$Q_{\text{cal}} = -(-54\underline{3}.4\ \text{J} + 2\underline{9}2\ \text{J})$$

$$Q_{\text{cal}} \cong +2\underline{5}1 \tag{4}$$

The calorimeter heat capacity is calculated from Equation 15.1:

$$C_{\text{cal}} = \frac{Q_{\text{cal}}}{\Delta t} = \frac{+2\underline{5}1\ \text{J}}{(27.0\ {}^{\circ}\text{C} - 20.0\ {}^{\circ}\text{C})}$$

$$C_{\text{cal}} \cong 3\underline{5}.8\ \text{J/}{}^{\circ}\text{C} \cong 36\ \text{J/}{}^{\circ}\text{C}$$

Using a calorimeter with a known heat capacity, the heat capacity or specific heat of an unknown sample can be easily determined.

EXAMPLE 15-4

Experimentally Determining Specific Heat

A calorimeter with a heat capacity of 36.0 J/°C contains 20.0 g of water at 50.0 °C. A 10.0 g metal alloy sample is at 20.0 °C. When the sample is placed into the calorimeter, the final equilibrium temperature is 49.0 °C. What is the specific heat of the alloy?

Solution:

The system is the calorimeter, the alloy sample, and the hot water. Equation 15.2 permits you to calculate the specific heat of the alloy.

$$c_{\text{sp alloy}} = \frac{Q_{\text{alloy}}}{(m\Delta t)_{\text{alloy}}} \qquad (1)$$

You need to determine the heat transfer of the alloy sample when it is placed into the calorimeter in order to solve Equation (1). From the conservation of energy principle, you know that

$$Q_{\text{hw}} + Q_{\text{alloy}} + Q_{\text{cal}} = 0 \text{ J}. \qquad (2)$$

First, determine the heat lost by the hot water:

$$Q_{\text{hw}} = (mc_{\text{sp}}\Delta t)_{\text{hw}} = [mc_{\text{sp}}(t_{\text{f}} - t_{\text{i}})]_{\text{hw}}$$

$$Q_{\text{hw}} = (20.0 \text{ g})(4.18 \text{ J/g}\cdot{}^\circ\text{C})(49.0\ {}^\circ\text{C} - 50.0\ {}^\circ\text{C})$$

$$Q_{\text{hw}} \cong -8\underline{3}.6 \text{ J}$$

Next, determine the heat lost by the calorimeter:

$$Q_{\text{cal}} = (C\Delta t)_{\text{cal}} = [C(t_{\text{f}} - t_{\text{i}})]_{\text{cal}}$$

$$Q_{\text{cal}} = (36.0 \text{ J/}{}^\circ\text{C})(49.0\ {}^\circ\text{C} - 50.0\ {}^\circ\text{C})$$

$$Q_{\text{cal}} \cong -3\underline{6}.0 \text{ J}$$

Solve Equation (2) for the heat gained by the alloy (Q_{alloy}):

$$Q_{\text{alloy}} = -(Q_{\text{hw}} + Q_{\text{cal}}) \qquad (3)$$

Substitute Equation (3) for Q_{alloy} into Equation (1) and determine $c_{\text{sp alloy}}$:

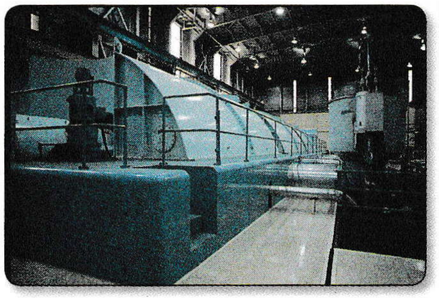
$$c_{\text{sp alloy}} = \frac{-(Q_{\text{hw}} + Q_{\text{cal}})}{(m\Delta t)_{\text{alloy}}} = \frac{-(Q_{\text{hw}} + Q_{\text{cal}})}{[m(t_{\text{f}} - t_{\text{i}})]_{\text{alloy}}}$$

$$c_{\text{sp alloy}} = \frac{-[(-8\underline{3}.6 \text{ J}) + (-3\underline{6}.0 \text{ J})]}{(10.0 \text{ g})(49.0\ {}^\circ\text{C} - 20.0\ {}^\circ\text{C})}$$

$$c_{\text{sp alloy}} \cong 0.4\underline{1}2 \text{ J/g}\cdot{}^\circ\text{C} \cong 0.41 \text{ J/g}\cdot{}^\circ\text{C}$$

Problem-Solving Strategy 15.2

Solving heat-balance problems requires finding the heat exchanged (both magnitude and sign) for every part of the system within the adiabatic boundary. The conservation of heat within an adiabatic boundary assures you that the sum of all heat transfers will be zero.



15-6 Engineers use calorimetric procedures in order to determine the rate of heat transfer and the efficiency of their power plants.

The procedure just described in this example is similar to many such heat balance calculations performed by scientists and engineers for a variety of reasons. For example, the latent heats required to bring about various phase transitions are determined using calorimetric procedures.

15.12 Heat and Phase Transitions

Every material has a characteristic specific heat capacity. You would expect that any thermal energy gained by an object would cause an increase in temperature, but this is not always the case.

Consider an ice cube in a freezer ice cube tray. Assume that we have inserted a thermocouple junction (an electrical temperature sensor) in the water before it freezes. When we are ready for the demonstration, we quickly remove the instrumented ice cube from the tray, connect it to a temperature display instrument, and immerse the cube in a beaker of ice and water. A laboratory burner is lighted under the beaker.

The initial temperature of the ice cube is around –10 °C, the temperature of the freezer. We record the temperature every ten seconds as the ice begins to warm. At first, the temperature rises rapidly. At 0 °C, however, we note that the cube's temperature stops rising. The ice cubes in the beaker, including our "wired" cube, are all shrinking in size as they melt. When the ice fragments are nearly gone, we observe that the temperature starts rising again, but more slowly than before. The water's temperature continues to rise at a steady rate until the water nears the boiling point (numerous small bubbles form and detach from the bottom of the beaker). Then the rate of temperature rise drops off, and the temperature becomes constant when the water reaches a rolling boil.

You may wonder why the temperature does not rise continuously during the entire period. The flame has been adding thermal energy to the beaker at the same rate the whole time. The answer lies in the water's changes of state. The thermal energy that the water absorbed without changing temperature was used to break the bonds between the water molecules in the ice. The same energy is released when the phase change is reversed. That is, the amount of thermal energy absorbed when ice melts is released when water freezes. Similarly, the amount of thermal energy absorbed when water vaporizes is released when steam condenses. This amount of thermal energy per unit mass is a constant for a given phase change of each substance. The **latent** (hidden) **heat of fusion (L_f)** of a substance is the amount of thermal energy required to melt 1 kg of the substance at its melting point. The **latent heat of vaporization (L_v)** is the amount of thermal energy required to vaporize 1 kg of a substance at its boiling point. The values of these quantities are determined at standard conditions so that they may be compared. Table 15-3 on page 349 lists some heats of fusion and vaporization for selected substances.

Heats of fusion and vaporization are determined by the following relationships:

$$L_f = \frac{Q_{\text{melt}}}{m} \tag{15.3}$$

$$L_v = \frac{Q_{\text{boil}}}{m} \tag{15.4}$$

Units of L_f and L_v are kJ/kg (or J/g).

TABLE 15-3

Heats of Fusion and Vaporization with Associated Phase Transition Temperatures for Selected Substances

Substance	Melting Point t_m (°C)	Heat of Fusion L_f (kJ/kg)	Boiling Point t_b (°C)	Heat of Vaporization L_v (kJ/kg)
Acetone	−94.8	98.0	56.1	501
Aluminum	660.3	397.0	2520	10 900
Carbon Dioxide	−56.6	205	—	—
Copper	1084.6	208.7	2563	4726
Ethanol	−114.1	78.4	78.3	837
Gold	1064.2	63.7	2857	1645
Iron	1538	247.3	—	—
Lead	327.5	23.0	1750	866
Mercury	−38.8	1.4	356.7	295
Nitrogen	−210.0	25.3	−195.8	198
Oxygen	−218.8	13.8	−183.0	213
Silver	961.8	104.7	2163	2323
Tungsten	3422	284.5	5900	4820
Water	0.0	333.5	100.0	2256

EXAMPLE 15-5

Determining Heat of Fusion: A Single Phase Change

When 10.0 g of ice is placed in a calorimeter with a heat capacity of 36.0 J/°C, the equilibrium temperature is 0.0 °C. Then 20.0 g of water at 50.0 °C is added, and the new equilibrium temperature is 5.2 °C. The ice has melted. Calculate the heat of fusion (L_f) of ice.

Solution:

The heat of fusion is calculated from Equation 15.3:

$$L_f = \frac{Q_{melt}}{m} \text{ or}$$

$$Q_{melt} = mL_f \qquad (1)$$

Note which quantities are changing and which are constant. The ice melts, but its temperature does not change. The cold meltwater from the ice (subscript "cw") warms from 0.0 °C to 5.2 °C. The warm water (subscript "ww") cools from 50.0 °C to 5.2 °C. The calorimeter experiences the same temperature change as the cold water. Therefore, heat flows from the warm water into the ice to melt it and into the resulting cold water and the calorimeter.

From the conservation of energy in an adiabatic container you know that

$$Q_{cal} + Q_{cw} + Q_{ww} + Q_{melt} = 0 \text{ J}.$$

Solve for Q_{melt}:

$$Q_{\text{melt}} = -(Q_{\text{cal}} + Q_{\text{cw}} + Q_{\text{ww}})$$

Substitute Equation (1) for Q_{melt} and solve for L_f:

$$m_{\text{ice}}L_f = -(Q_{\text{cal}} + Q_{\text{cw}} + Q_{\text{ww}})$$

$$L_f = \frac{-(Q_{\text{cal}} + Q_{\text{cw}} + Q_{\text{ww}})}{m_{\text{ice}}} \tag{2}$$

Now calculate each of the heat expressions in Equation (2):

Calorimeter: $Q_{\text{cal}} = (C\Delta t)_{\text{cal}} = (36.0\ \text{J/°C})(+5.2\ \text{°C})$

$Q_{\text{cal}} \cong +1\underline{8}7\ \text{J}$

Cold water: $Q_{\text{cw}} = (mc_{\text{sp}}\Delta t)_{\text{cw}} = (10.0\ \text{g})(4.18\ \text{J/g·°C})(+5.2\ \text{°C})$

$Q_{\text{cw}} \cong +2\underline{1}7\ \text{J}$

Warm water: $Q_{\text{ww}} = (mc_{\text{sp}}\Delta t)_{\text{ww}} = (20.0\ \text{g})(4.18\ \text{J/g·°C})(-44.8\ \text{°C})$

$Q_{\text{ww}} \cong -37\underline{4}5\ \text{J}$

Solve Equation (2) for L_f:

$$L_f = \frac{-[1\underline{8}7\ \text{J} + 2\underline{1}7\ \text{J} + (-37\underline{4}5\ \text{J})]}{10.0\ \text{g}}$$

$$L_f \cong 3\underline{3}4\ \text{J/g} \cong 330\ \text{J/g}$$

This result, obtained in a typical high school laboratory calorimeter, agrees reasonably well with the accepted value of 333.5 J/g for the heat of fusion of water.

EXAMPLE 15-6

From Ice to Steam: Thermal Energy Needed for Two Phase Changes

How much thermal energy is required to raise the temperature of 10.0 g of water from −10.0 °C to 110.0 °C?

Solution:

This problem involves five parts. We must determine the thermal energy changes for the following:

(1) Heating the ice from −10.0 °C to 0.0 °C (Q_{ice});

(2) Converting the ice to liquid water at 0.0 °C (Q_{melt});

(3) Heating the liquid water from 0.0 °C to its boiling point at 100.0 °C (Q_{w});

(4) Converting the liquid water to gaseous water (steam) at 100.0 °C (Q_{boil});

(5) Heating the steam from 100.0 °C to 110.0 °C (Q_{s}).

$$Q_{\text{total}} = Q_{\text{ice}} + Q_{\text{melt}} + Q_{\text{w}} + Q_{\text{boil}} + Q_{\text{s}} \tag{1}$$

Heats of fusion and vaporization can be obtained from Table 15-3. Specific heats can be obtained from Table 15-1.

Determine the heat transferred for each term on the right side of Equation (1):

Heat ice:

$$Q_{\text{ice}} = (mc_{\text{sp}}\Delta t)_{\text{ice}} = (10.0\ \text{g})(2.09\ \text{J/g}\cdot{}^\circ\text{C})(+10.0\ {}^\circ\text{C})$$
$$Q_{\text{ice}} = +20\underline{9}.0\ \text{J}$$

Melt ice:

$$Q_{\text{melt}} = m_{\text{ice}}L_{\text{f}} = (10.0\ \text{g})(333.5\ \text{J/g})$$
$$Q_{\text{melt}} = +33\underline{3}5\ \text{J}$$

Heat water:

$$Q_{\text{w}} = (mc_{\text{sp}}\Delta t)_{\text{w}} = (10.0\ \text{g})(4.18\ \text{J/g}\cdot{}^\circ\text{C})(+100.0\ {}^\circ\text{C})$$
$$Q_{\text{w}} = +41\underline{8}0\ \text{J}$$

Boil water:

$$Q_{\text{boil}} = m_{\text{w}}L_{\text{v}} = (10.0\ \text{g})(2256\ \text{J/g})$$
$$Q_{\text{boil}} = +22\,\underline{5}60\ \text{J}$$

Heat steam:

$$Q_{\text{s}} = (mc_{\text{sp}}\Delta t)_{\text{s}} = (10.0\ \text{g})(2.01\ \text{J/g}\cdot{}^\circ\text{C})(+10.0\ {}^\circ\text{C})$$
$$Q_{\text{s}} = +20\underline{1}.0\ \text{J}$$

Sum all of the heat quantities to find the total thermal energy transferred:

$$Q_{\text{total}} = 20\underline{9}\ \text{J} + 33\underline{3}5\ \text{J} + 41\underline{8}0\ \text{J} + 22\,\underline{5}60\ \text{J} + 20\underline{1}\ \text{J}$$
$$Q_{\text{total}} \cong 30\,\underline{4}85\ \text{J}$$
$$Q_{\text{total}} \cong 30\,500\ \text{J (Allowable precision is to nearest 100 J.)}$$

Problem-Solving Strategy 15.3

Calculating the thermal energy required to change a solid substance to a vapor, or vice versa, can be broken down into five heat-transfer steps. The key to success is to use the proper specific heat value for the corresponding phase and the correct latent heat for each phase change. Be sure to observe proper significant digit rules when summing the final result.

15B Section Review

1. Discuss why, when evaluating heat processes involving water, it might be more convenient to use the calorie unit.
2. Define and clearly differentiate the terms temperature, heat, and thermal energy.
3. Discuss the difference between the heat capacity of a solid steel cylinder and its specific heat.
4. As thermal energy is added to a solid, why does its average temperature stop rising when it begins to melt?
5. What is the quantity of heat required to melt a gram of a substance at its melting point called? What do we call the quantity of heat per gram that must be removed to freeze the substance at its freezing point?
6. In a laboratory exercise, you are required to tentatively identify a 65.0 g sample of metal from Table 15-1 by its specific heat, using a calorimeter. The calorimeter and the 50.0 g of distilled water that it contains are both at a temperature of 24.0 °C. The initial temperature of the metal just before it is placed into the calorimeter is 98.5 °C. The final temperature of the calorimeter and its contents is 26.5 °C. The calorimeter's heat capacity is 39.5 J/°C. What metal is the sample most likely made of?
7. How much heat must be lost in order to completely condense 100.0 g of nitrogen at its condensation point?

15B Objectives

After completing this section, I can

- ✓ define the heat capacity of an object.
- ✓ compare specific heat capacity with heat capacity and explain why the former is more descriptive of a substance.
- ✓ using the principle of the conservation of thermal energy, determine heat capacities for various objects and materials.
- ✓ explain why solids have heats of fusion and liquids have heats of fusion and vaporization.
- ✓ perform heat balance calculations involving phase transitions.