

Chapter 14 Temperature and Heat

"The determination of temperature has long been recognized as a problem of the greatest importance in physical science. It has accordingly been made a subject of most careful attention, and, especially in late years, of very elaborate and refined experimental researches; and we are thus at present in possession of as complete a practical solution of the problem as can be desired, even for the most accurate investigation." William Thompson, Lord Kelvin

14.1 Temperature

The simplest and most intuitive definition of temperature is that *temperature is a measure of the hotness or coldness of a body*. That is, if a body is hot it has a high temperature, if it is cold it has a low temperature. This is not a very good definition, as we will see in a moment, but it is one that most people have a "feel" for, because we all know what hot and cold is. Or do we?

Let us reconsider the "thought experiment" treated in chapter 1. We place three beakers on the table, as shown in figure 14.1. Several ice cubes are placed into the first beaker of water, whereas boiling water is poured into the third beaker. We place equal amounts of the ice water from beaker one and the boiling water from beaker three into the second beaker to form a mixture. I now take my left hand and plunge it into beaker one, and conclude that

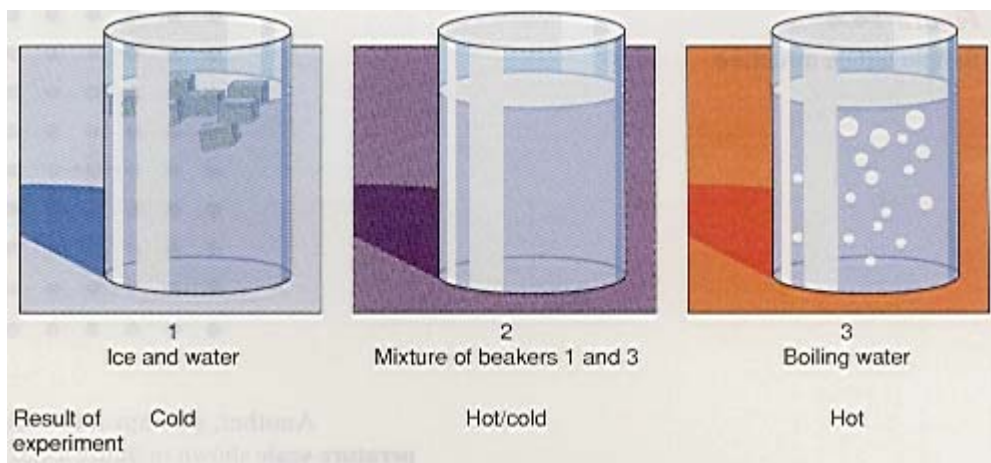


Figure 14.1 A "thought experiment" on temperature.

it is *cold*. After drying off my left hand, I place it into the middle mixture. After coming from the ice water, the mixture in the second beaker feels hot by comparison. So I conclude that the mixture is *hot*.

I now take my right hand and plunge it into the boiling water of beaker three. (This is of course the reason why this is only a "thought experiment.") I conclude that the water in beaker three is certainly *hot*. Drying off my hand again I then place it into beaker two. After the boiling water, the mixture feels cold by comparison, so I conclude that the mixture is *cold*. After this relatively scientific experiment, my conclusion is contradictory. That is, I found the middle mixture to be either hot or cold depending on the sequence of the measurement. Thus, the hotness or coldness of a body is not a good concept to use to define the temperature of a body. Although we may have an intuitive feel for hotness or coldness, we can not use our intuition for any precise scientific work.

The Thermometer

In order to make a measurement of the temperature of a body, a new technique, other than estimating hotness or coldness, must be found. Let us look for some characteristic of matter that changes as it is heated. The simplest such characteristic is that most materials expand when they are heated. Using this characteristic of matter we take a glass tube and fill it with a liquid, as shown in figure 14.2. When the liquid is heated it expands and rises up the tube. The height of the liquid in the tube can be used to measure the hotness or coldness of a body. The device will become a thermometer.

In order to quantify the process, we need to place numerical values on the glass tube, thus assigning a number that can be associated with the hotness or coldness of a body. This is the process of *calibrating* the thermometer.

First, we place the thermometer into the mixture of ice and water of beaker 1 in figure 14.1. The liquid lowers to a certain height in the glass tube. We scratch a mark on the glass at that height, and arbitrarily call it 0 degrees. Since it is the point where ice is melting in the water, we call 0° the melting point of ice. (Or similarly, the freezing point of water.)

Then we place the glass tube into beaker three, which contains the boiling water. (We assume that heat is continuously applied to beaker three to keep the water boiling.) The liquid in the glass tube is thus heated and expands to a new height. We mark this new height on the glass tube and arbitrarily call it 100° . Since the water is boiling at this point, we call it the boiling point of water.

Because the liquid in the tube expands linearly, to a first approximation, the distance between 0° and 100° can be divided into 100 equal parts. Any one of these divisions can be further divided into fractions of a degree. Thus, we obtain a complete scale of temperatures ranging from 0 to 100 degrees. Then we place this thermometer into the mixture of beaker two. The liquid in the glass rises to some number, and that number, whatever it may be, is the temperature of the mixture. That number is a numerical measure of the hotness or coldness of the body. We call this device a **thermometer**, and in particular *this scale of temperature that has 0° for the melting point of ice and 100° for the boiling point of water is called the **Celsius temperature scale** and is shown in figure 14.3(a). This scale is named after the Swedish astronomer, Anders Celsius, who proposed it in 1742.*

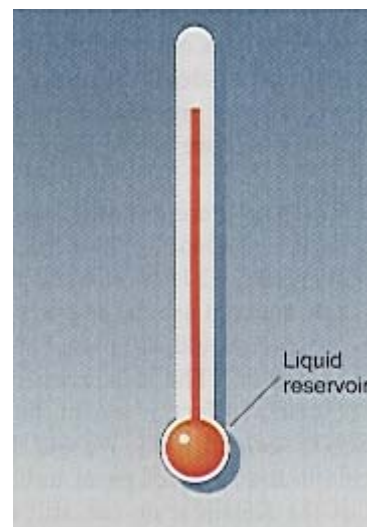


Figure 14.2 A thermometer.

Another, perhaps more familiar, temperature scale is the **Fahrenheit temperature scale** shown in figure 14.3(b). The melting point of ice on this scale is 32°F and the boiling point of water is 212°F . At first glance

it might seem rather strange to pick 32° for the freezing point and 212° for the boiling point of water. As a matter of fact Gabriel Fahrenheit, the German physicist, was not trying to use pure water as his calibration points. When the scale was first made, 0°F corresponded to the lowest temperature then known, the temperature of freezing brine (a salt water mixture), and 100°F was meant to be the temperature of the human body. Fahrenheit proposed his scale in 1714.

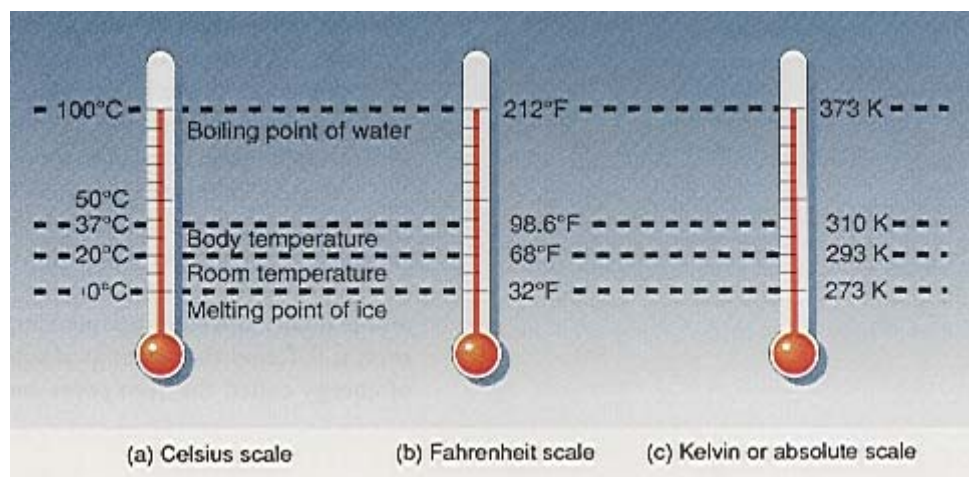


Figure 14.3 The temperature scales.

In addition to the Celsius and Fahrenheit scales there are other temperature scales, the most important of which is the Kelvin or absolute scale, as shown in figure 14.3(c). The melting point of ice on this scale is 273 K and the boiling point of water is 373 K. The **Kelvin temperature scale** does not use the degree symbol for a temperature. To use the terminology correctly, we should say that, “zero degrees Celsius corresponds to a temperature of 273 Kelvin.” The Kelvin scale is extremely important in dealing with the behavior of gases. In fact, it was in the study of gases that Lord Kelvin first proposed the absolute scale in 1848. We will discuss this more natural introduction to the Kelvin scale in the study of gases in chapter 15. For the present, however, the implications of the Kelvin scale can still be appreciated by looking at the molecular structure of a solid.

The simplest picture of a solid, if it could be magnified trillions of times, is a large array of atoms or molecules in what is called a lattice structure, as shown in figure 14.4. Each dot in the figure represents an atom or molecule, depending on the nature of the substance. Each molecule is in equilibrium with all the molecules around it. The molecule above exerts a force upward on the molecule, whereas the molecule below exerts a force downward. Similarly, there are balanced forces from right and left and in and out. The molecule is therefore in equilibrium. In fact every molecule of the solid is in equilibrium. When heat is applied to a solid body, the added energy causes a molecule to vibrate around its equilibrium position. As any one molecule vibrates, it interacts with its nearest neighbors causing them to vibrate, which in turn causes its nearest neighbors to vibrate, and so on.

Hence, the heat energy applied to the solid shows up as vibrational energy of the molecules of the solid. The higher the temperature of the solid, the larger is the vibrational motion of its molecules. The lower the temperature, the smaller is the vibrational motion of its molecules. Thus, the temperature of a body is really a measure of the mean or average kinetic energy of the vibrating molecules of the body.

It is therefore conceivable that if you could lower and lower the temperature of the body, the motion of the molecules would become less and less until at some very low temperature, the vibrational motion of the molecules would cease altogether. They would be frozen in one position. This point is called *absolute zero*, and is 0 on the Kelvin temperature scale. From work in quantum

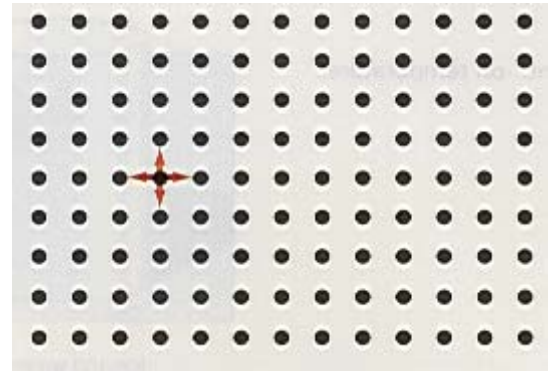


Figure 14.4 Simple lattice structure.

mechanics, however, it is found that even at absolute zero, the molecules contain a certain amount of energy called the *zero point energy*.

Even though temperature is really a measure of the mean kinetic energy of the molecules of a substance, from an experimental point of view it is difficult to make a standard of temperature in this way. Therefore, the International System of units considers temperature to be a first fundamental quantity and it is added to the four fundamental quantities of length, mass, time, and electric charge. *The SI unit of temperature is the kelvin, and is defined as 1/273.16 of the temperature of the triple point of water.* The triple point of water is that point on a pressure-temperature diagram where the three phases of water, the solid, the liquid, and the gas, can coexist in equilibrium at the same pressure and temperature.

Temperature Conversions

The Celsius temperature scale is the recognized temperature scale in most scientific work and in most countries of the world. The Fahrenheit scale will eventually become obsolete along with the entire British engineering system of units. For the present, however, it is still necessary to convert from one temperature scale to another. That is, if a temperature is given in degrees Fahrenheit, how can it be expressed in degrees Celsius, and vice versa? It is easy to see how this conversion can be made.

The principle of the thermometer is based on the linear expansion of the liquid in the tube. For two identical glass tubes containing the same liquid, the expansion of the liquid is the same in both tubes. Therefore, the height of the liquid columns is the same for each thermometer, as shown in figure 14.5. The ratio of these heights in each thermometer is also equal. Therefore,

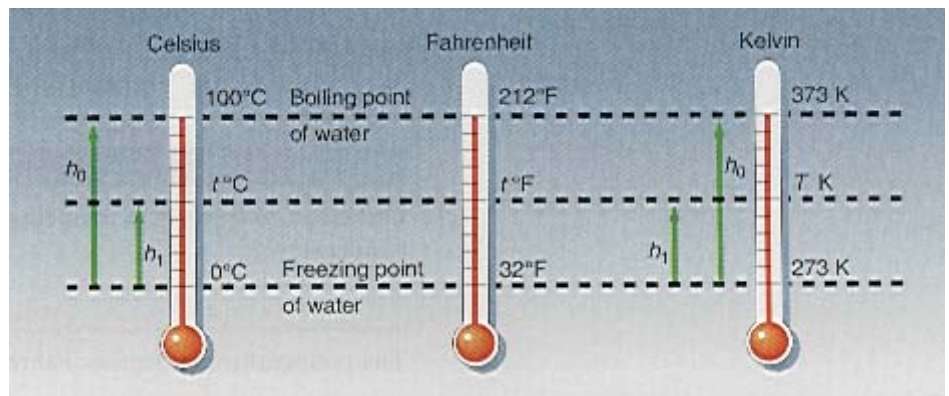


Figure 14.5 Converting one temperature scale to another.

$$\left(\frac{h_1}{h_0}\right)_{\text{Celsius}} = \left(\frac{h_1}{h_0}\right)_{\text{Fahrenheit}}$$

These ratios, found from figure 14.5, are

$$\frac{t^{\circ}\text{C} - 0^{\circ}}{100^{\circ} - 0^{\circ}} = \frac{t^{\circ}\text{F} - 32^{\circ}}{212^{\circ} - 32^{\circ}}$$

$$\frac{t^{\circ}\text{C}}{100^{\circ}} = \frac{t^{\circ}\text{F} - 32^{\circ}}{180^{\circ}}$$

Solving for the temperature in degrees Celsius

$$t \text{ } ^\circ\text{C} = \frac{100^\circ}{180^\circ}(t \text{ } ^\circ\text{F} - 32^\circ)$$

Simplifying,

$$t \text{ } ^\circ\text{C} = \frac{5}{9}(t \text{ } ^\circ\text{F} - 32^\circ) \quad (14.1)$$

Equation 14.1 allows us to convert a temperature in degrees Fahrenheit to degrees Celsius.

Example 14.1

Fahrenheit to Celsius. If room temperature is 68°F , what is this temperature in Celsius degrees?

Solution

The temperature in Celsius degrees, found from equation 14.1, is

$$\begin{aligned} t \text{ } ^\circ\text{C} &= \frac{5}{9}(t \text{ } ^\circ\text{F} - 32^\circ) = \frac{5}{9}(68^\circ - 32^\circ) = \frac{5}{9}(36) \\ &= 20 \text{ } ^\circ\text{C} \end{aligned}$$

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To convert a temperature in degrees Celsius to one in Fahrenheit, we solve equation 14.1 for $t \text{ } ^\circ\text{F}$ to obtain

$$t \text{ } ^\circ\text{F} = \frac{9}{5}t \text{ } ^\circ\text{C} + 32^\circ \quad (14.2)$$

Example 14.2

Celsius to Fahrenheit. A temperature of -5.00°C is equivalent to what Fahrenheit temperature?

Solution

The temperature in degrees Fahrenheit, found from equation 14.2, is

$$\begin{aligned} t \text{ } ^\circ\text{F} &= \frac{9}{5}t \text{ } ^\circ\text{C} + 32^\circ = \frac{9}{5}(-5.00^\circ) + 32^\circ = -9 + 32^\circ \\ &= 23 \text{ } ^\circ\text{F} \end{aligned}$$

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We can also find a conversion of absolute temperature to Celsius temperatures from figure 14.5, as

$$\begin{aligned} \left(\frac{h_1}{h_0}\right)_{\text{Celsius}} &= \left(\frac{h_1}{h_0}\right)_{\text{Kelvin}} \\ \frac{t \text{ } ^\circ\text{C} - 0^\circ}{100^\circ - 0^\circ} &= \frac{T \text{ K} - 273}{373 - 273} \end{aligned}$$

$$\frac{t\text{ }^{\circ}\text{C}}{100} = \frac{T\text{ K} - 273}{100}$$

Therefore, the conversion of Kelvin temperature to Celsius temperatures is given by

$$t\text{ }^{\circ}\text{C} = T\text{ K} - 273 \quad (14.3)$$

And the reverse conversion by

$$T\text{ K} = t\text{ }^{\circ}\text{C} + 273 \quad (14.4)$$

For very precise work, 0 °C is actually equal to 273.16 K. In such cases, equations 14.3 and 14.4 should be modified accordingly.

Example 14.3

Celsius to Kelvin. Normal room temperature is considered to be 20.0 °C, find the value of this temperature on the Kelvin scale.

Solution

The absolute temperature, found from equation 14.4, is

$$T\text{ K} = t\text{ }^{\circ}\text{C} + 273 = 20.0 + 273 = 293\text{ K}$$

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Note, in this book we will try to use the following convention: temperatures in Celsius and Fahrenheit will be represented by the lower case t , whereas Kelvin or absolute temperatures will be represented by a capital T . However, in some cases where time and temperature are found in the same equation, the lower case t will be used for time, and the upper case T will be used for temperature regardless of the unit used for temperature.

14.2 Heat

A solid body is composed of trillions upon trillions of atoms or molecules arranged in a lattice structure, as shown in figure 14.4. Each of these molecules possess an electrical potential energy and a vibrational kinetic energy. *The sum of the potential energy and kinetic energy of all these molecules is called the **internal energy of the body**.* When that internal energy is transferred between two bodies as a result of the difference in temperatures between the two bodies it is called heat.

Heat is thus the amount of internal energy flowing from a body at a higher temperature to a body at a lower temperature. Hence, a body does not contain heat, it contains internal energy. When the body cools, its internal energy is decreased; when it is heated, its internal energy is increased. A useful analogy is to compare the internal energy of a body to the money you have in a savings bank, whereas heat is analogous to the deposits or withdrawals of money.

Whenever two bodies at different temperatures are brought into contact, thermal energy always flows from the hotter body to the cooler body until they are both at the same temperature. When this occurs we say the two bodies are in **thermal equilibrium**. This is essentially the principle behind the thermometer. The thermometer is placed in contact with the body whose temperature is desired. Thermal energy flows from the hotter body to the cooler body until thermal equilibrium is reached. At that point, the thermometer is at the same temperature as the body. Hence, the thermometer is capable of measuring the temperature of a body.

The traditional unit of heat was the *kilocalorie*, which was defined as the quantity of heat required to raise the temperature of 1 kg of water 1 °C, from 14.5 °C to 15.5 °C. It may seem strange to use the unit of kilocalorie for heat since heat is a flow of energy, and the unit of energy is a joule. Historically it was not known that heat was a form of energy, but rather it was assumed that heat was a material quantity contained in bodies and was called Caloric. It was assumed that a hot body contained a great deal of caloric while a cold body

contained only a small quantity of caloric. It was not until Benjamin Thompson's (1753-1814) experiments on the boring of cannons in 1798, that it became known that heat was, in fact, a form of energy. Later James Prescott Joule (1818-1889) performed experiments to show *the exact equivalence between mechanical energy and heat energy*. That equivalence is called the **mechanical equivalent of heat** and is

$$1 \text{ kilocalorie} = 1000 \text{ calories} = 4186 \text{ J}$$

The unit of heat in the British engineering system is the **British thermal unit**, abbreviated Btu. *One Btu is the heat required to raise the temperature of 1 lb of water 1 °F, from 58.5 °F to 59.5 °F*. The relation between the Btu, the kilocalorie (kcal), the foot-pound (ft lb), and the joule is

$$1 \text{ Btu} = 0.252 \text{ kcal} = 778 \text{ ft lb} = 1055 \text{ J}$$

In terms of the SI unit of energy, the joule, it takes 4186 J of energy to raise the temperature of 1 kg of water 1 °C, from 14.5 °C to 15.5 °C.

We should also mention that the kilocalorie is sometimes called the large calorie and is identical to the unit used by dietitians. Thus when dietitians specify a diet as consisting of 1500 calories a day, they really mean that it is 1500 kcal per day.

14.3 Specific Heat

When the temperature of several substances is raised the same amount, each substance does not absorb the same amount of thermal energy. This can be shown by Tyndall's demonstration in figure 14.6.

Four balls made of aluminum, iron, brass, and lead, all of the same mass, are placed in a beaker of boiling water, as shown in figure 14.6(a). After about 10 or 15 minutes, these balls will reach thermal equilibrium with the water and will all be at the same temperature as the boiling water. The four balls are then placed on a piece of paraffin, as shown in figure 14.6(b). Almost immediately, the aluminum ball melts the wax and falls through the paraffin, as shown in figure

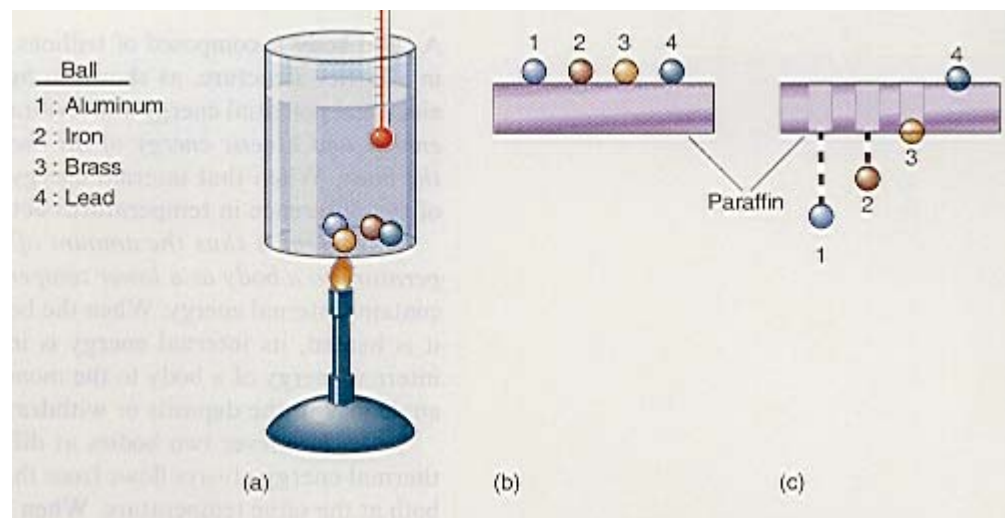


Figure 14.6 Tyndall's demonstration.

14.6(c). A little later in time the iron ball melts its way through the wax. The brass ball melts part of the wax and sinks into it deeply. However, it does not melt enough wax to fall through. The lead ball barely melts the wax and sits on the top of the sheet of paraffin.

How can this strange behavior of the four different balls be explained? Since each ball was initially in the boiling water, each absorbed energy from the boiling water. When the balls were placed on the sheet of paraffin, each ball gave up that energy to the wax, thereby melting the wax. But since each ball melted a different amount of wax in a given time, each ball must have given up a different amount of energy to the wax. Therefore each ball must have absorbed a different quantity of energy while it was in the boiling water. Hence, *different bodies absorb a different quantity of thermal energy even when subjected to the same temperature change*.

To handle the problem of different bodies absorbing different quantities of thermal energy when subjected to the same temperature change, the **specific heat** c of a body is defined as the amount of thermal energy Q required to raise the temperature of a unit mass of the material 1 °C. In terms of the SI unit joules, the specific heat c of a body is defined as the number of joules Q required to raise the temperature of 1 kg of the material 1 °C. Thus,

$$c = \frac{Q}{m\Delta t} \quad (14.5)$$

We observe from this definition that the specific heat of water in SI units is 4186 J/kg °C, since 4186 J raises the temperature of 1 kg of water 1 °C. All other materials have a different value for the specific heat. Some specific heats are shown in table 14.1. Note that water has the largest specific heat.

Having defined the specific heat by equation 14.5, we can rearrange that equation into the form

$$Q = mc\Delta t \quad (14.6)$$

Equation 14.6 represents the amount of thermal energy Q that will be absorbed or liberated in any process.

Using equation 14.6 it is now easier to explain the Tyndall demonstration. The thermal energy absorbed by each ball while in the boiling water is

$$\begin{aligned} Q_{\text{Al}} &= mc_{\text{Al}}\Delta t \\ Q_{\text{iron}} &= mc_{\text{iron}}\Delta t \\ Q_{\text{brass}} &= mc_{\text{brass}}\Delta t \\ Q_{\text{lead}} &= mc_{\text{lead}}\Delta t \end{aligned}$$

Because all the balls went from room temperature to 100 °C, the boiling point of water, they all experienced the same temperature change Δt . Because all the masses were equal, the thermal energy absorbed by each ball is directly proportional to its specific heat. We can observe from table 14.1 that

$$\begin{aligned} c_{\text{Al}} &= 900 \text{ J/(kg } ^\circ\text{C)} \\ c_{\text{iron}} &= 452 \text{ J/(kg } ^\circ\text{C)} \\ c_{\text{brass}} &= 394 \text{ J/(kg } ^\circ\text{C)} \\ c_{\text{lead}} &= 130 \text{ J/(kg } ^\circ\text{C)} \end{aligned}$$

Because the specific heat of aluminum is the largest of the four materials, the aluminum ball absorbs the greatest amount of thermal energy while in the water. Hence, it also liberates the greatest amount of thermal energy to melt the wax and should be the first ball to melt through the wax. Iron, brass, and lead absorb less thermal energy respectively because of their lower specific heats and consequently liberate thermal energy to melt the wax in this same sequence. Hence, Tyndall's demonstration can be explained by the different specific heats of the four materials.

If the masses are not the same, then the amount of thermal energy absorbed depends on the product of the mass m and the specific heat c . The ball with the largest value of mc absorbs the most heat energy.

Example 14.4

Absorption of thermal energy. A steel ball at room temperature is placed in a pan of boiling water. If the mass of the ball is 200 g, how much thermal energy is absorbed by the ball?

Solution

The thermal energy absorbed by the ball, given by equation 14.6, is

$$\begin{aligned} Q &= mc\Delta t \\ &= (0.200 \text{ kg}) \left(452 \frac{\text{J}}{\text{kg } ^\circ\text{C}} \right) (100 ^\circ\text{C} - 20.0 ^\circ\text{C}) \\ &= 7230 \text{ J} \end{aligned}$$

Table 14.1 Specific Heats of Various Materials	
Material	$\frac{\text{J}}{\text{kg } ^\circ\text{C}}$
Air	1009
Aluminum	900.0
Brass	393.5
Copper	385.1
Glass	837.2
Gold	129.8
Iron	452.1
Lead	129.8
Platinum	134.0
Silver	238.6
Steel	452.1
Tin	226.0
Tungsten	134.0
Zinc	389.3
Water	4186
Ice	2093
Steam	2013

An interesting thing to note is that once the ball reaches the 100 °C mark, it is at the same temperature as the water and hence, there is no longer a transfer of thermal energy into the ball no matter how long the ball is left in the boiling water. All the thermal energy supplied to the pot containing the ball and the water will then go into boiling away the water.

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Example 14.5

The final temperature. If a 500-g aluminum block at an initial temperature of 10.0 °C absorbs 85500 J of energy in a thermal process, what will its new temperature be?

Solution

The specific heat of aluminum, found from table 14.1, is 900 J/(kg °C). The change in temperature is found as

$$Q = mc\Delta t$$

$$\Delta t = \frac{Q}{mc}$$

$$\Delta t = \frac{(85500 \text{ J})}{(0.5 \text{ kg})(900 \text{ J/kg } ^\circ\text{C})}$$

$$\Delta t = 190 \text{ } ^\circ\text{C}$$

The final temperature is found from

$$\Delta t = t_f - t_i$$

and hence

$$t_f = \Delta t + t_i$$

$$t_f = (190 \text{ } ^\circ\text{C}) + (10 \text{ } ^\circ\text{C})$$

$$t_f = 200 \text{ } ^\circ\text{C}$$

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14.4 Calorimetry

Calorimetry is defined as the measurement of heat. These measurements are performed in a device called a **calorimeter**. The simplest of all calorimeters consists of a metal container placed on a plastic insulating ring inside a larger highly polished metallic container, as shown in figure 14.7. The space between the two containers is filled with air to minimize the thermal energy lost from the inner calorimeter cup to the environment. The highly polished outer container reflects any external radiated energy that might otherwise make its way to the inner cup. A plastic cover is placed on the top of the calorimeter to prevent any additional loss of thermal energy to the environment. The inner cup is thus insulated from the environment, and all measurements of thermal energy absorption or liberation are made here. A thermometer is placed through a hole in the cover so that the temperature inside the calorimeter can be measured. The calorimeter is used to measure the specific heat of various substances, and the latent heat of fusion and vaporization of water.

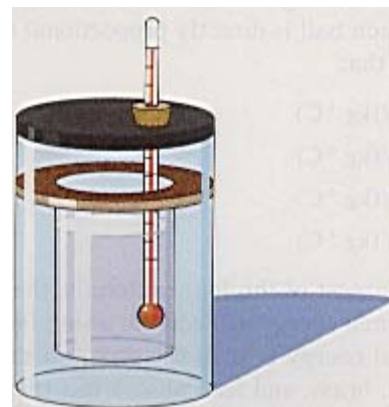


Figure 14.7 A calorimeter.

The basic principle underlying the calorimeter is the conservation of energy. The thermal energy lost by those bodies that lose thermal energy is equal to the thermal energy gained by those bodies that gain thermal energy. We write this conservation principle mathematically as

$$\text{Thermal energy lost} = \text{Thermal energy gained} \quad (14.7)$$

As an example of the use of the calorimeter, let us determine the specific heat of a sample of iron of mass m_s . We place the iron sample in a pot of boiling water until the iron sample eventually reaches the temperature of boiling water, namely 100 °C. Meanwhile we place the inner calorimeter cup on a scale and determine its mass m_c . Then we place water within the cup and again place it on the scale to determine its mass. The difference between these two scale readings is the mass of water m_w in the cup. We place the inner cup into the calorimeter and place a thermometer through a hole in the cover of the calorimeter so that the initial temperature of the water t_{iw} is measured.

After the iron sample reaches 100 °C, we place it within the inner calorimeter cup, and close the cover quickly. As time progresses, the temperature of the water, as recorded by the thermometer, starts to rise. It eventually stops at a final equilibrium temperature t_{fw} of the water, the sample, and the calorimeter can. The iron sample was the hot body and it lost thermal energy, whereas the water and the can, which is in contact with the water, absorb this thermal energy as is seen by the increased temperature of the mixture. We analyze the problem by the conservation of energy, equation 14.7, as

$$\begin{aligned} \text{Thermal energy lost} &= \text{Thermal energy gained} \\ Q_s &= Q_w + Q_c \end{aligned} \quad (14.8)$$

That is, the thermal energy lost by the sample Q_s is equal to the thermal energy gained by the water Q_w plus the thermal energy gained by the calorimeter cup Q_c . However, the thermal energy absorbed or liberated in any process, given by equation 14.6, is

$$Q = mc\Delta t$$

Using equation 14.6 in equation 14.8, gives

$$m_s c_s \Delta t_s = m_w c_w \Delta t_w + m_c c_c \Delta t_c \quad (14.9)$$

where

- m_s is the mass of the sample
- m_w is the mass of the water
- m_c is the mass of the calorimeter cup
- c_s is the specific heat of the sample
- c_w is the specific heat of the water
- c_c is the specific heat of the calorimeter cup

The change in the temperature of the sample is the difference between its initial temperature of 100 °C and its final equilibrium temperature t_{fw} . That is,

$$\Delta t_s = 100 \text{ }^\circ\text{C} - t_{fw} \quad (14.10)$$

The change in temperature of the water and calorimeter cup are equal since the water is in contact with the cup and thus has the same temperature. Therefore,

$$\Delta t_w = \Delta t_c = t_{fw} - t_{iw} \quad (14.11)$$

Substituting equations 14.10 and 14.11 into equation 14.9, yields

$$m_s c_s (100 - t_{fw}) = m_w c_w (t_{fw} - t_{iw}) + m_c c_c (t_{fw} - t_{iw}) \quad (14.12)$$

All the quantities in equation 14.12 are known except for the specific heat of the sample, c_s . Solving for the specific heat yields

$$c_s = \frac{m_w c_w (t_{fw} - t_{iw}) + m_c c_c (t_{fw} - t_{iw})}{m_s (100 - t_{fw})} \quad (14.13)$$

Example 14.6

Find the specific heat. A 0.0700-kg iron specimen is used to determine the specific heat of iron. The following laboratory data were found:

$$\begin{aligned}m_s &= 0.0700 \text{ kg} & t_{iw} &= 20.0 \text{ }^\circ\text{C} \\m_c &= 0.0600 \text{ kg} & t_{fw} &= 23.5 \text{ }^\circ\text{C} \\c_c &= 900 \text{ J/kg }^\circ\text{C} & m_w &= 0.150 \text{ kg} \\& & t_s &= 100 \text{ }^\circ\text{C}\end{aligned}$$

Find the specific heat of the specimen.

Solution

The specific heat of the iron specimen, found from equation 14.13, is

$$\begin{aligned}c_s &= \frac{m_w c_w (t_{fw} - t_{iw}) + m_c c_c (t_{fw} - t_{iw})}{m_s (100 - t_{fw})} \\&= \frac{(0.150 \text{ kg})(4186 \text{ J/kg }^\circ\text{C})(23.5 \text{ }^\circ\text{C} - 20.0 \text{ }^\circ\text{C})}{(0.0700 \text{ kg})(100 \text{ }^\circ\text{C} - 23.5 \text{ }^\circ\text{C})} \\&\quad + \frac{(0.0600 \text{ kg})(900 \text{ J/kg }^\circ\text{C})(23.5 \text{ }^\circ\text{C} - 20.0 \text{ }^\circ\text{C})}{(0.0700 \text{ kg})(100 \text{ }^\circ\text{C} - 23.5 \text{ }^\circ\text{C})} \\&= 446 \text{ J/kg }^\circ\text{C}\end{aligned}$$

which is in good agreement with the accepted value of the specific heat of iron of 452 J/kg $^\circ\text{C}$

[To go to this Interactive Example click on this sentence.](#)

14.5 Change of Phase

Matter exists in three states called the **phases of matter**. They are the solid phase, the liquid phase, and the gaseous phase. Let us see how one phase of matter is changed into another.

Let us examine the behavior of matter when it is heated over a relatively large range of temperatures. In particular, let us start with a piece of ice at $-20.0 \text{ }^\circ\text{C}$ and heat it to a temperature of $120 \text{ }^\circ\text{C}$. We place the ice inside a strong, tightly sealed, windowed enclosure containing a thermometer. Then we apply heat, as shown in figure 14.8. We observe the temperature as a function of time and plot it, as in figure 14.9.

As the heat is applied to the solid ice, the temperature of the block increases with time until $0 \text{ }^\circ\text{C}$ is reached. At this point the temperature remains constant, even though heat is being continuously applied. Looking at the block of ice, through the window in the container, we observe small drops of liquid water forming on the block of ice. The ice is starting to melt. We observe that the temperature remains constant until every bit of the solid ice is converted into the liquid water. We are observing a **change of phase**. That is, the ice is changing from the solid phase into the liquid phase. As soon as all the ice is melted, we again

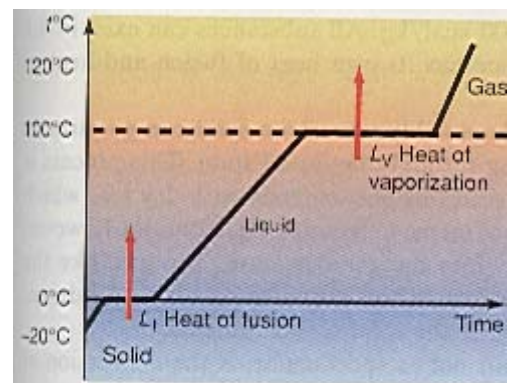
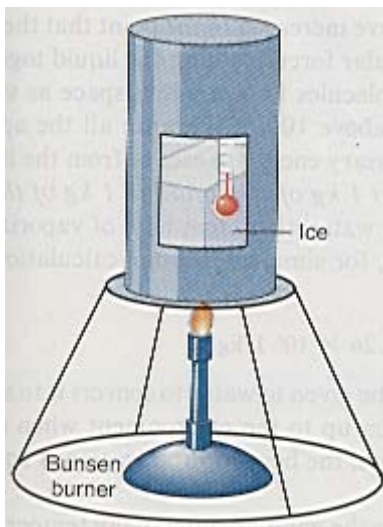


Figure 14.8 Converting ice to water to steam. **Figure 14.9** Changes of phase.

observe an increase in the temperature of the liquid water. The temperature increases up to 100 °C, and then levels off. Thermal energy is being applied, but the temperature is not changing. Looking through the window into the container, we see that there are bubbles forming throughout the liquid. The water is boiling. The liquid water is being converted to steam, the gaseous state of water. The temperature remains at this constant value of 100 °C until every drop of the liquid water has been converted to the gaseous steam. After that, as we continuously supply heat, we observe an increase in the temperature of the steam. Superheated steam is being made. (Note, you should not try to do this experiment on your own, because enormous pressures can be built up by the steam, causing the closed container to explode.)

Let us go back and analyze this experiment more carefully. As the thermal energy was supplied to the below freezing ice, its temperature increased to 0 °C. At this point the temperature remained constant even though heat was being continuously applied. Where did this thermal energy go if the temperature never changed? The thermal energy went into the melting of the ice, changing its phase from the solid to the liquid phase. If we observe the solid in terms of its lattice structure, figure 14.4, we can see that each molecule is vibrating about its equilibrium position. As heat is applied, the vibration increases, until at 0 °C, the vibrations of the molecules become so intense that the molecules literally pull apart from one another changing the entire structure of the material. This is the melting process. The amount of heat necessary to tear these molecules apart is a constant and is called the latent heat of fusion of that material. *The **latent heat of fusion** L_f , is the amount of heat necessary to convert 1 kg of the solid to 1 kg of the liquid.* For water, it is found experimentally that it takes 334,000 J of thermal energy to melt 1 kg of ice. Hence we take the latent heat of fusion of water to be

$$L_f = 3.34 \times 10^5 \text{ J/kg}$$

If we must supply $3.34 \times 10^5 \text{ J/kg}$ to melt ice, then we must take away $3.34 \times 10^5 \text{ J/kg}$ to freeze water. Thus, the heat of fusion is equal to the heat of melting. The word *latent* means hidden or invisible, and not detectable as a temperature change. Heat supplied that does change the temperature is called *sensible heat*, because it can be sensed by a thermometer.

In the liquid state there are still molecular forces holding the molecules together, but because of the energy and motion of the molecules, these forces can not hold the molecules in the relatively rigid position they had in the solid state. This is why the liquid is able to flow and take the shape of any container in which it is placed.

As the water at 0 °C is further heated, the molecules absorb more and more energy, increasing their mean velocity within the liquid. This appears as a rise in temperature of the liquid. At 100 °C, so much energy has been imparted to the water molecules, that the molecular speeds have increased to the point that the molecules are ready to pull away from the molecular forces holding the liquid together. As further thermal energy is applied, the molecules fly away into space as steam. The temperature of the water does not rise above 100 °C because all the applied heat is supplying the molecules with the necessary energy to escape from the liquid.

*The heat that is necessary to convert 1 kg of the liquid to 1 kg of the gas is called the **latent heat of vaporization** L_v .* For water, it is found experimentally that it takes 2,260,000 J of thermal energy to boil 1 kg of liquid water. Hence we take the latent heat of vaporization of water to be

$$L_v = 2.26 \times 10^6 \text{ J/kg}$$

Because this amount of thermal energy must be given to water to convert it to steam, this same quantity of thermal energy is given up to the environment when steam condenses back into the liquid state. Therefore, the heat of vaporization is equal to the heat of condensation.

Liquid water can also be converted to the gaseous state at any temperature, a process called *evaporation*. Thus, water left in an open saucer overnight will be gone by morning. Even though the temperature of the water remained at the room temperature, the liquid was converted to a gas. It evaporated into the air. The gaseous state of water is then usually referred to as *water vapor* rather than steam. At 0 °C the latent heat of vaporization is $2.51 \times 10^6 \text{ J/kg}$. All substances can exist in the three states of matter, and each substance has its own heat of fusion and heat of vaporization.

Note also that another process is possible whereby a solid can go directly to a gas and vice versa without ever going through the liquid state. This process is called **sublimation**. Many students have seen this phenomenon with dry ice (which is carbon dioxide in the solid state). The ice seems to be smoking. Actually, however, the solid carbon dioxide is going directly into the gaseous state. The gas, like the dry ice, is so cold that it causes water vapor in the surrounding air to condense, which is seen as the “smoky” clouds around the solid carbon dioxide.

A more common phenomena, but not as spectacular, is the conversion of water vapor, a gas, directly into ice crystals, a solid, in the sublimation process commonly known as frost. On wintry mornings when you first get up and go outside your home, you sometimes see ice all over the tips of the grass in the yard and over the windshield and other parts of your car. The water vapor in the air did not first condense to water droplets and then the water droplets froze. Instead, the grass and the car surfaces were so cold that the water vapor in the air went directly from the gaseous state into the solid state without ever going through the liquid state.

The reverse process whereby the solid goes directly into the gas also occurs in nature, but it is not as noticeable as frost. There are times in the winter when a light covering of snow is observed on the ground. The temperature may remain below freezing, and an overcast sky may prevent any sun from heating up or melting that snow. Yet, in a day or so, some of that snow will have disappeared. It did not melt, because the temperature always remained below freezing. Some of the snow crystals went directly into the gaseous state as water vapor.

Just as there is a latent heat of fusion L_f and latent heat of vaporization L_v there is also a latent heat of sublimation L_s . Its value is given by

$$L_s = 2.83 \times 10^6 \text{ J/kg}$$

Thus, the heat that is necessary to convert 1.00 kg of the solid ice into 1.00 kg of the gaseous water vapor is called the **latent heat of sublimation**.

It is interesting to note here that there is no essential difference in the water molecule when it is either a solid, a liquid, or a gas. The molecule consists of the same two hydrogen atoms bonded to one oxygen atom. The difference in the state is related to the different energy, and hence speed of the molecule in the different states.

Notice that it takes much more energy to convert 1 kg of water to 1 kg of steam, than it does to convert 1 kg of ice to 1 kg of liquid water, almost seven times as much. This is also why a steam burn can be so serious, since the steam contains so much energy. Let us now consider some more examples.

Example 14.7

Converting ice to steam. Let us compute the thermal energy that is necessary to convert 5.00 kg of ice at -20.0°C to superheated steam at 120°C .

Solution

The necessary thermal energy is given by

$$Q = Q_i + Q_f + Q_w + Q_v + Q_s \quad (14.14)$$

where

Q_i is the energy needed to heat the ice up to 0°C

Q_f is the energy needed to melt the ice

Q_w is the energy needed to heat the water to 100°C

Q_v is the energy needed to boil the water

Q_s is the energy needed to heat the steam to 120°C

The necessary thermal energy to warm up the ice from -20.0°C to 0°C is found from

$$Q_i = m_i c_i [0^\circ - (-20.0^\circ\text{C})]$$

The latent heat of fusion is the amount of heat needed per kilogram to melt the ice. The total amount of heat needed to melt all the ice is the heat of fusion times the number of kilograms of ice present. Hence, the thermal energy needed to melt the ice is

$$Q_f = m_i L_f \quad (14.15)$$

The thermal energy needed to warm the water from 0°C to 100°C is

$$Q_w = m_w c_w (100^\circ\text{C} - 0^\circ\text{C})$$

The latent heat of vaporization is the amount of heat needed per kilogram to boil the water. The total amount of heat needed to boil all the water is the heat of vaporization times the number of kilograms of water present. Hence, the thermal energy needed to convert the liquid water at 100°C to steam at 100°C is

$$Q_v = m_w L_v \quad (14.16)$$

and

$$Q_s = m_s c_s (120\text{ }^\circ\text{C} - 100\text{ }^\circ\text{C})$$

is the thermal energy needed to convert the steam at $100\text{ }^\circ\text{C}$ to superheated steam at $120\text{ }^\circ\text{C}$. Substituting all these equations into equation 14.14 gives

$$Q = m_i c_i [0\text{ }^\circ\text{C} - (-20\text{ }^\circ\text{C})] + m_i L_f + m_w c_w (100\text{ }^\circ\text{C} - 0\text{ }^\circ\text{C}) + m_w L_v + m_s c_s (120\text{ }^\circ\text{C} - 100\text{ }^\circ\text{C}) \quad (14.17)$$

Using the values of the specific heat from table 14.1, we get

$$\begin{aligned} Q &= (5.00\text{ kg}) \left(2093 \frac{\text{J}}{\text{kg } ^\circ\text{C}} \right) (20.0\text{ }^\circ\text{C}) + (5.00\text{ kg}) \left(3.34 \times 10^5 \frac{\text{J}}{\text{kg}} \right) \\ &+ (5.00\text{ kg}) \left(4186 \frac{\text{J}}{\text{kg } ^\circ\text{C}} \right) (100.0\text{ }^\circ\text{C}) + (5.00\text{ kg}) \left(2.26 \times 10^6 \frac{\text{J}}{\text{kg}} \right) \\ &+ (5.00\text{ kg}) \left(2013 \frac{\text{J}}{\text{kg } ^\circ\text{C}} \right) (20.0\text{ }^\circ\text{C}) \\ &= 0.209 \times 10^6\text{ J} + 1.67 \times 10^6\text{ J} + 2.09 \times 10^6\text{ J} + 11.3 \times 10^6\text{ J} + 0.201 \times 10^6\text{ J} \\ &= 15.5 \times 10^6\text{ J} \end{aligned}$$

Therefore, we need $15.5 \times 10^6\text{ J}$ of thermal energy to convert 5.00 kg of ice at $-20.0\text{ }^\circ\text{C}$ to superheated steam at $120\text{ }^\circ\text{C}$. Note the relative size of each term's contribution to the total thermal energy.

[To go to this Interactive Example click on this sentence.](#)

Example 14.8

Latent heat of fusion. The heat of fusion of water L_f can be found in the laboratory using a calorimeter. If 31.0 g of ice m_i at $0\text{ }^\circ\text{C}$ are placed in a 60.0-g calorimeter cup m_c that contains 170 g of water m_w at an initial temperature t_{iw} of $20.0\text{ }^\circ\text{C}$, after the ice melts, the final temperature of the water t_{fw} is found to be $5.57\text{ }^\circ\text{C}$. Find the heat of fusion of water from this data. The specific heat of the calorimeter is $900\text{ J/kg } ^\circ\text{C}$.

Solution

From the fundamental principle of calorimetry

$$\begin{aligned} \text{Thermal energy gained} &= \text{Thermal energy lost} \\ Q_f + Q_{iw} &= Q_w + Q_c \end{aligned} \quad (14.18)$$

where Q_f is the thermal energy necessary to melt the ice through the fusion process and Q_{iw} is the thermal energy necessary to warm the water that came from the melted ice. We call this water ice water to distinguish it from the original water in the container. This liquid water is formed at $0\text{ }^\circ\text{C}$ and will be warmed to the final equilibrium temperature of the mixture t_{fw} . The thermal energy lost by the original water in the calorimeter is Q_w , and Q_c is the thermal energy lost by the calorimeter itself. Equation 14.18 therefore becomes

$$m_i L_f + m_{iw} c_w (t_{fw} - 0\text{ }^\circ\text{C}) = m_w c_w (t_{iw} - t_{fw}) + m_c c_c (t_{iw} - t_{fw})$$

We find the heat of fusion by solving for L_f , as

$$L_f = \frac{(m_w c_w + m_c c_c)(t_{iw} - t_{fw}) - m_{iw} c_w (t_{fw} - 0 \text{ } ^\circ\text{C})}{m_i} \quad (14.19)$$

Since the laboratory data were taken in grams we convert them to kilograms and the heat of fusion is found as

$$L_f = \frac{[(0.170 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C}) + (0.060 \text{ kg})(900 \text{ J/kg } ^\circ\text{C})](20.0 \text{ } ^\circ\text{C} - 5.57 \text{ } ^\circ\text{C}) - (0.031 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C})(5.57 \text{ } ^\circ\text{C} - 0 \text{ } ^\circ\text{C})}{0.031 \text{ kg}}$$

$$L_f = 3.33 \times 10^5 \text{ J/kg}$$

Note that this is in very good agreement with the standard value of $3.34 \times 10^5 \text{ J/kg}$.

[To go to this Interactive Example click on this sentence.](#)

Example 14.9

Latent heat of vaporization. The heat of vaporization L_v of water can be found in the laboratory by passing steam at $100 \text{ } ^\circ\text{C}$ into a calorimeter containing water. As the steam condenses and cools it gives up thermal energy to the water and the calorimeter. In the experiment the following data were taken:

mass of calorimeter cup	$m_c = 60.0 \text{ g}$
mass of water	$m_w = 170 \text{ g}$
mass of condensed steam	$m_s = 3.00 \text{ g}$
initial temperature of water	$t_{iw} = 19.9 \text{ } ^\circ\text{C}$
final temperature of water	$t_{fw} = 30.0 \text{ } ^\circ\text{C}$
specific heat of calorimeter	$c_c = 900 \text{ J/kg } ^\circ\text{C}$

Find the heat of vaporization from this data.

Solution

To determine the heat of vaporization let us start with the fundamental principle of calorimetry

$$\begin{aligned} \text{Thermal energy lost} &= \text{Thermal energy gained} \\ Q_v + Q_{sw} &= Q_w + Q_c \end{aligned} \quad (14.20)$$

where Q_v is the thermal energy necessary to condense the steam and Q_{sw} is the thermal energy necessary to cool the water that came from the condensed steam. We use the subscript sw to remind us that this is the water that came from the steam in order to distinguish it from the original water in the container. This liquid water is formed at $100 \text{ } ^\circ\text{C}$ and will be cooled to the final equilibrium temperature of the mixture t_{fw} . Here Q_w is the thermal energy gained by the original water in the calorimeter and Q_c is the thermal energy gained by the calorimeter itself. Equation 14.20 therefore becomes

$$m_s L_v + m_{sw} c_w (100 \text{ } ^\circ\text{C} - t_{fw}) = m_w c_w (t_{fw} - t_{iw}) + m_c c_c (t_{fw} - t_{iw})$$

Solving for the heat of vaporization,

$$L_v = \frac{m_w c_w (t_{fw} - t_{iw}) + m_c c_c (t_{fw} - t_{iw}) - m_{sw} c_w (100 \text{ } ^\circ\text{C} - t_{fw})}{m_s} \quad (14.21)$$

Therefore,

$$L_v = \frac{(0.170 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C})(30.0 \text{ } ^\circ\text{C} - 19.9 \text{ } ^\circ\text{C}) + (0.060 \text{ kg})(900 \text{ J/kg } ^\circ\text{C})(30.0 - 19.9 \text{ } ^\circ\text{C}) - (0.003 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C})(100 \text{ } ^\circ\text{C} - 30.0)}{0.003 \text{ kg}}$$

Thus, we find from the experimental data that the heat of vaporization is

$$L_v = 2.28 \times 10^6 \text{ J/kg}$$

which is in good agreement with the standard value of $2.26 \times 10^6 \text{ J/kg}$.

[To go to this Interactive Example click on this sentence.](#)

Example 14.10

Mixing ice and water. If 10.0 g of ice, at 0 °C, are mixed with 50.0 g of water at 80.0 °C, what is the final temperature of the mixture?

Solution

When the ice is mixed with the water it will gain heat from the water. The law of conservation of thermal energy becomes

$$\begin{aligned} \text{Thermal energy gained} &= \text{Thermal energy lost} \\ Q_f + Q_{iw} &= Q_w \end{aligned}$$

where Q_f is the heat gained by the ice as it goes through the melting process. When the ice melts, it becomes water at 0 °C. Let us call this water ice water to distinguish it from the original water in the container. Thus, Q_{iw} is the heat gained by the ice water as it warms up from 0 °C to the final equilibrium temperature t_{fw} . Finally, Q_w is the heat lost by the original water, which is at the initial temperature t_{iw} . Thus,

$$m_i L_f + m_{iw} c_w (t_{fw} - 0 \text{ } ^\circ\text{C}) = m_w c_w (t_{iw} - t_{fw})$$

where m_i is the mass of the ice, m_{iw} is the mass of the ice water, and m_w is the mass of the original water. Solving for the final temperature of the water we get

$$\begin{aligned} m_i L_f + m_{iw} c_w t_{fw} &= m_w c_w t_{iw} - m_w c_w t_{fw} \\ m_{iw} c_w t_{fw} + m_w c_w t_{fw} &= m_w c_w t_{iw} - m_i L_f \\ (m_{iw} c_w + m_w c_w) t_{fw} &= m_w c_w t_{iw} - m_i L_f \\ t_{fw} &= \frac{m_w c_w t_{iw} - m_i L_f}{m_{iw} c_w + m_w c_w} \end{aligned}$$

The final equilibrium temperature of the water becomes

$$\begin{aligned} t_{fw} &= \frac{(0.050 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C})(80.0 \text{ } ^\circ\text{C}) - (0.010 \text{ kg})(3.35 \times 10^5 \text{ J/kg})}{(0.010 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C}) + (0.050 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C})} \\ &= \frac{16744 \text{ J} - 3350 \text{ J}}{251 \text{ J/} ^\circ\text{C}} \\ &= 53.3 \text{ } ^\circ\text{C} \end{aligned}$$

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Example 14.11

Something is wrong here. Repeat example 14.10 with the initial temperature of the water at 10.0 °C.

Solution

Using the same equation as for the final water temperature in example 14.10, we get

$$t_{fw} = \frac{m_w c_w t_{iw} - m_i L_f}{m_{iw} c_w + m_w c_w}$$

Thus,

$$\begin{aligned} t_{fw} &= \frac{(0.050 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C})(10.0 \text{ } ^\circ\text{C}) - (0.010 \text{ kg})(3.35 \times 10^5 \text{ J/kg})}{(0.010 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C}) + (0.050 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C})} \\ &= \frac{2093 \text{ J} - 3350 \text{ J}}{251 \text{ J/}^\circ\text{C}} \\ &= -5.00 \text{ } ^\circ\text{C} \end{aligned}$$

There is something very wrong here! Our answer says that the final temperature is 5° below zero. But this is impossible. The temperature of the water can not be below 0 °C and still be water, and the ice that was placed in the water can not convert all the water to ice and cause all the ice to be at a temperature of 5° below zero. Something is wrong. Let us check our equation. The equation worked for the last example, why not now? The equation was derived with the assumption that all the ice that was placed in the water melted. Is this a correct assumption? The energy necessary to melt all the ice is found from

$$Q_f = m_i L_f = (0.01 \text{ kg})(3.35 \times 10^5 \text{ J/kg}) = 3350 \text{ J}$$

The energy available to melt the ice comes from the water. The maximum thermal energy available occurs when all the water is cooled to 0 °C. Therefore, the maximum available energy is

$$\begin{aligned} Q_w &= m_w c_w (t_{iw} - 0 \text{ } ^\circ\text{C}) = (0.05 \text{ kg})(4186 \text{ J/kg } ^\circ\text{C})(10.0 \text{ } ^\circ\text{C}) \\ &= 2093 \text{ J} \end{aligned}$$

The amount of energy available to melt all the ice is 2093 J and it would take 3350 J to melt all the ice present. Therefore, there is not enough energy to melt the ice. Hence, our initial assumption that all the ice melted is incorrect. Thus, our equation is no longer valid. There is an important lesson to be learned here. All through our study of physics we make assumptions in order to derive equations. If the assumptions are correct, the equations are valid and can be used to predict some physical phenomenon. If the assumptions are not correct, the final equations are useless. In this problem there is still ice left and hence the final temperature of the mixture is 0 °C. The amount of ice that actually melted can be found by using the relation

$$f Q_f = Q_w$$

where f is the fraction of the ice that melts. Thus,

$$\begin{aligned} f &= \frac{Q_w}{Q_f} \\ &= \frac{2093 \text{ J}}{3350 \text{ J}} \\ &= 0.625 \end{aligned}$$

Therefore, only 62.5% of the ice melted and the final temperature of the mixture is 0 °C.

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The Language of Physics

Temperature

The simplest definition of temperature is that temperature is a measure of the hotness or coldness of a body. A better definition is that temperature is a measure of the mean kinetic energy of the molecules of the body (p.).

Thermometer

A device for measuring the temperature of a body (p.).

Celsius temperature scale

A temperature scale that uses 0° for the melting point of ice and 100° for the boiling point of water (p.).

Fahrenheit temperature scale

A temperature scale that uses 32° for the melting point of ice and 212° for the boiling point of water (p.).

Kelvin temperature scale

The absolute temperature scale. The lowest temperature attainable is absolute zero, the 0 K of this scale. The temperature for the melting point of ice is 273 K and 373 K for the boiling point of water (p.).

Internal energy

The sum of the potential and kinetic energy of all the molecules of a body (p.).

Heat

The flow of thermal energy from a body at a higher temperature to a body at a lower

temperature. When a body cools, its internal energy is decreased; when it is heated, its internal energy is increased (p.).

Thermal equilibrium

Whenever two bodies at different temperatures are touched together, thermal energy always flows from the hotter body to the cooler body until they are both at the same temperature. When this occurs the two bodies are said to be in thermal equilibrium (p.).

Kilocalorie

An older unit of heat. It is defined as the amount of thermal energy required to raise the temperature of 1 kg of water 1°C (p.).

British thermal unit (Btu)

The unit of heat in the British engineering system of units. It is the amount of thermal energy required to raise the temperature of 1 lb of water 1°F (p.).

Mechanical equivalent of heat

The equivalence between mechanical energy and thermal energy. One kcal is equal to 4186 J (p.).

Specific heat

A characteristic of a material. It is defined as the number of joules of energy required to raise the temperature of 1 kg of the material 1°C (p.). The specific heat of water is $4186 \text{ J/kg }^{\circ}\text{C}$.

Calorimetry

The measurement of heat (p.).

Calorimeter

An instrument that is used to make measurements of heat. The basic principle underlying the calorimeter is the conservation of energy. The thermal energy lost by those bodies that lose thermal energy is equal to the thermal energy gained by those bodies that gain thermal energy (p.).

Phases of matter

Matter exists in three phases, the solid phase, the liquid phase, and the gaseous phase (p.).

Change of phase

The change in a body from one phase of matter to another. As an example, melting is a change from the solid state of a body to the liquid state. Boiling is a change in state from the liquid state to the gaseous state (p.).

Latent heat of fusion

The amount of heat necessary to convert 1 kg of the solid to 1 kg of the liquid (p.).

Latent heat of vaporization

The amount of heat necessary to convert 1 kg of the liquid to 1 kg of the gas (p.).

Summary of Important Equations

Convert Fahrenheit temperature to Celsius $t^{\circ}\text{C} = \frac{5}{9}(t^{\circ}\text{F} - 32^{\circ})$ (14.1)

Convert Celsius temperature to Fahrenheit $t^{\circ}\text{F} = \frac{9}{5}t^{\circ}\text{C} + 32^{\circ}$ (14.2)

Convert Celsius temperature to Kelvin $T^{\circ}\text{K} = t^{\circ}\text{C} + 273$ (14.4)

Thermal energy absorbed or liberated $Q = mc\Delta t$ (14.6)

Principle of calorimetry
Thermal energy gained
= Thermal energy lost (14.7)

Fusion $Q_f = m_i L_f$ (14.15)

Vaporization $Q_v = m_w L_v$ (14.16)

Questions for Chapter 14

1. What is the difference between temperature and heat?

2. Explain how a bathtub of water at 5°C can contain more thermal energy than a cup of coffee at 95°C .

3. Discuss how the human body uses the latent heat of vaporization to cool itself through the process of evaporation.

*4. Relative humidity is defined as the percentage of the amount of water vapor in the air to the

maximum amount of water vapor that the air can hold at that temperature. Discuss how the relative humidity affects the process of evaporation in general and how it affects the human body in particular.

*5. It is possible for a gas to go directly to the solid state without going through the liquid state, and vice versa. The process is called *sublimation*. An example of such a process is the formation of frost.

Discuss the entire process of sublimation, the latent heat involved, and give some more examples of the process.

*6. Why does ice melt when an object is placed upon it? Describe the process of ice skating from the pressure of the skate on the ice.

Problems for Chapter 14

14.1 Temperature

1. Convert the following normal body temperatures to degrees Celsius: (a) oral temperature of 98.6°F , (b) rectal temperature of 99.6°F , and (c) axial (armpit) temperature of 97.6°F .

2. Find the value of absolute zero on the Fahrenheit scale.

3. For what value is the Fahrenheit temperature equal to the Celsius temperature?

4. Convert the following temperatures to Fahrenheit: (a) 38.0°C , (b) 68.0°C , (c) 250°C , (d) -10.0°C , and (e) -20.0°C .

5. Convert the following Fahrenheit temperatures to Celsius: (a) -23.0°F , (b) 12.5°F , (c) 55.0°F , (d) 90.0°F , and (e) 180°F .

6. A temperature change of 5°F corresponds to what temperature change in Celsius degrees?

*7. Derive an equation to convert the temperature in Fahrenheit degrees to its corresponding Kelvin temperature.

*8. Derive an equation to convert the change in temperature in Celsius degrees to a change in temperature in Fahrenheit degrees.

14.3 Specific Heat

9. A 450-g ball of copper at 20.0°C is placed in a pot of boiling water until equilibrium is reached.

How much thermal energy is absorbed by the ball?

10. A 250-g glass marble is taken from a freezer at -23.0°C and placed into a beaker of boiling water. How much thermal energy is absorbed by the marble?

11. How much thermal energy must be supplied by an electric immersion heater if you wish to raise the temperature of 5.00 kg of water from 20.0°C to 100°C ?

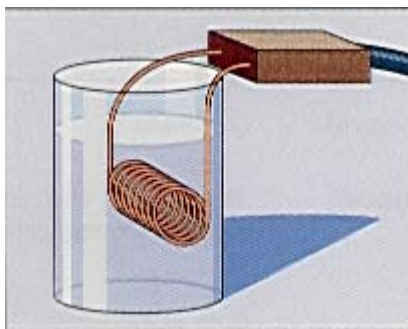


Diagram for problem 11.

12. A 2.00-kg mass of copper falls from a height of 3.00 m to an insulated floor. What is the maximum possible temperature increase of the copper?

13. An iron block slides down an iron inclined plane at a constant speed. The plane is 10.0 m long and is inclined at an angle of 35.0° with the horizontal. Assuming that half the energy lost to friction goes into the block, what is the difference in

temperature of the block from the top of the plane to the bottom of the plane?

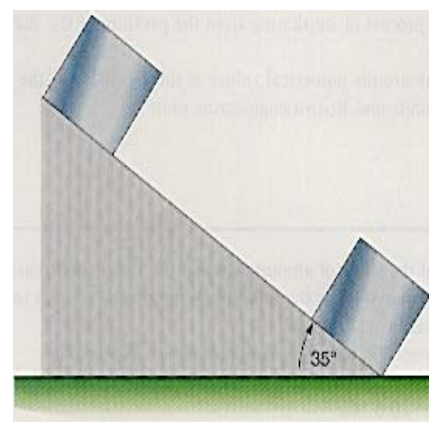


Diagram for problem 13.

14. A 2000-kg car is traveling at 96.6 km/hr when it is braked to a stop. What is the maximum possible thermal energy generated in the brakes?

15. How much thermal energy is absorbed by an aluminum ball 20.0 cm in diameter, initially at a temperature of 20.0°C , if it is placed in boiling water?

14.4 Calorimetry

16. If 30.0 g of water at 5.00°C are mixed with 50.0 g of water at 70.0°C and 25.0 g of water at 100°C , find the resultant temperature of the mixture.

17. If 80.0 g of lead shot at 100°C is placed into 100 g of water

at $20.0\text{ }^{\circ}\text{C}$ in an aluminum calorimeter of 60.0-g mass, what is the final temperature?

18. A 100-g mass of an unknown material at $100\text{ }^{\circ}\text{C}$ is placed in an aluminum calorimeter of 60.0 g that contains 150 g of water at an initial temperature of $20.0\text{ }^{\circ}\text{C}$. The final temperature is observed to be $21.5\text{ }^{\circ}\text{C}$. What is the specific heat of the substance and what substance do you think it is?

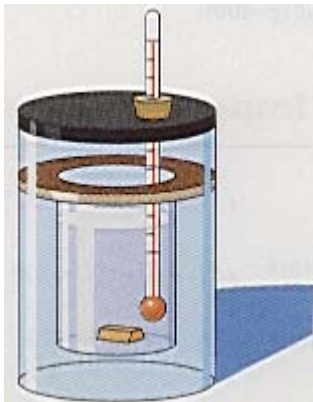


Diagram for problem 18.

19. A 100-g mass of an unknown material at $100\text{ }^{\circ}\text{C}$, is placed in an aluminum calorimeter of 60.0-g mass that contains 150 g of water at an initial temperature of $15.0\text{ }^{\circ}\text{C}$. At equilibrium the final temperature is $19.5\text{ }^{\circ}\text{C}$. What is the specific heat of the material and what material is it?

20. How much water at $50.0\text{ }^{\circ}\text{C}$ must be added to 60.0 kg of water at $10.0\text{ }^{\circ}\text{C}$ to bring the final mixture to $20.0\text{ }^{\circ}\text{C}$?

21. A 100-g aluminum calorimeter contains 200 g of water at $15.0\text{ }^{\circ}\text{C}$. If 100.0 g of lead at $50.0\text{ }^{\circ}\text{C}$ and 60.0 g of copper at $60.0\text{ }^{\circ}\text{C}$ are placed in the calorimeter, what is the final temperature in the calorimeter?

22. A 200-g piece of platinum is placed inside a furnace until it is in thermal equilibrium. The platinum is then placed in a 100-g aluminum calorimeter containing 400 g of water at $5.00\text{ }^{\circ}\text{C}$. If the final equilibrium temperature of the

water is $10.0\text{ }^{\circ}\text{C}$, find the temperature of the furnace.

14.5 Change of Phase

23. How many joules are needed to change 50.0 g of ice at $-10.0\text{ }^{\circ}\text{C}$ to water at $20.0\text{ }^{\circ}\text{C}$?

24. If 50.0 g of ice at $0.0\text{ }^{\circ}\text{C}$ are mixed with 50.0 g of water at $80.0\text{ }^{\circ}\text{C}$ what is the final temperature of the mixture?

25. How much ice at $0\text{ }^{\circ}\text{C}$ must be mixed with 50.0 g of water at $75.0\text{ }^{\circ}\text{C}$ to give a final water temperature of $20\text{ }^{\circ}\text{C}$?

26. If 50.0 g of ice at $0.0\text{ }^{\circ}\text{C}$ are mixed with 50.0 g of water at $20.0\text{ }^{\circ}\text{C}$, what is the final temperature of the mixture? How much ice is left in the mixture?

27. How much heat is required to convert 10.0 g of ice at $-15.0\text{ }^{\circ}\text{C}$ to steam at $105\text{ }^{\circ}\text{C}$?

28. In the laboratory, 31.0 g of ice at $0\text{ }^{\circ}\text{C}$ is placed into an 85.0-g copper calorimeter cup that contains 155 g of water at an initial temperature of $23.0\text{ }^{\circ}\text{C}$. After the ice melts, the final temperature of the water is found to be $6.25\text{ }^{\circ}\text{C}$. From this laboratory data, find the heat of fusion of water and the percentage error between the standard value and this experimental value.

29. A 100-g iron ball is heated to $100\text{ }^{\circ}\text{C}$ and then placed in a hole in a cake of ice at $0.00\text{ }^{\circ}\text{C}$. How much ice will melt?

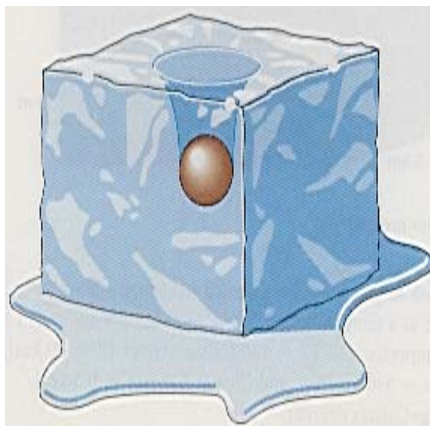


Diagram for problem 29.

30. How much steam at $100\text{ }^{\circ}\text{C}$ must be mixed with 300 g of water at $20.0\text{ }^{\circ}\text{C}$ to obtain a final water temperature of $80.0\text{ }^{\circ}\text{C}$?

31. How much steam at $100\text{ }^{\circ}\text{C}$ must be mixed with 1 kg of ice at $0.00\text{ }^{\circ}\text{C}$ to produce water at $20.0\text{ }^{\circ}\text{C}$?

32. In the laboratory, 6.00 g of steam at $100\text{ }^{\circ}\text{C}$ is placed into an 85.0-g copper calorimeter cup that contains 155 g of water at an initial temperature of $18.5\text{ }^{\circ}\text{C}$. After the steam condenses, the final temperature of the water is found to be $41.0\text{ }^{\circ}\text{C}$. From this laboratory data, find the heat of vaporization of water and the percentage error between the standard value and this experimental value.

*33. An electric stove is rated at 1 kW of power. If a pan containing 1.00 kg of water at $20.0\text{ }^{\circ}\text{C}$ is placed on this stove, how long will it take to boil away all the water?

34. An electric immersion heater is rated at 0.200 kW of power. How long will it take to boil 100 cm^3 of water at an initial temperature of $20.0\text{ }^{\circ}\text{C}$?

Additional Problems

35. A 890-N man consumes 3000 kcal of food per day. If this same energy were used to heat the same weight of water, by how much would the temperature of the water change?

36. An electric space heater is rated at 1.50 kW of power. How many kcal of thermal energy does it produce per second? How many Btu's of thermal energy per hour does it produce?

*37. A 0.055-kg mass of lead at an initial temperature of $135\text{ }^{\circ}\text{C}$, a 0.075-kg mass of brass at an initial temperature of $185\text{ }^{\circ}\text{C}$, and a 0.0445-kg of ice at an initial temperature of $-5.25\text{ }^{\circ}\text{C}$ is placed into a calorimeter containing 0.250 kg of water at an initial temperature of $23.0\text{ }^{\circ}\text{C}$. The aluminum calorimeter has a mass of 0.085 kg . Find the final temperature of the mixture.

38. A 100-g lead bullet is fired into a fixed block of wood at a speed of 350 m/s. If the bullet comes to rest in the block, what is the maximum change in temperature of the bullet?

*39. A 35-g lead bullet is fired into a 6.5-kg block of a ballistic pendulum that is initially at rest. The combined bullet-pendulum rises to a height of 0.125 m. Find (a) the speed of the combined bullet-pendulum after the collision, (b) the original speed of the bullet, (c) the original kinetic energy of the bullet, (d) the kinetic energy of the combined bullet-pendulum after the collision, and (e) how much of the initial mechanical energy was converted to thermal energy in the collision. If 50% of the energy lost shows up as thermal energy in the bullet, what is the change in energy of the bullet?

*40. After 50.0 g of ice at 0 °C is mixed with 200 g of water, also at 0 °C, in an insulated cup of 15.0-cm radius, a paddle wheel, 15.0 cm in radius, is placed inside the cup and set into rotational motion. What force, applied at the end of the paddle wheel, is necessary to rotate the paddle wheel at 60 rpm, for 10.0 minutes such that the final temperature of the mixture will be 15.0 °C?

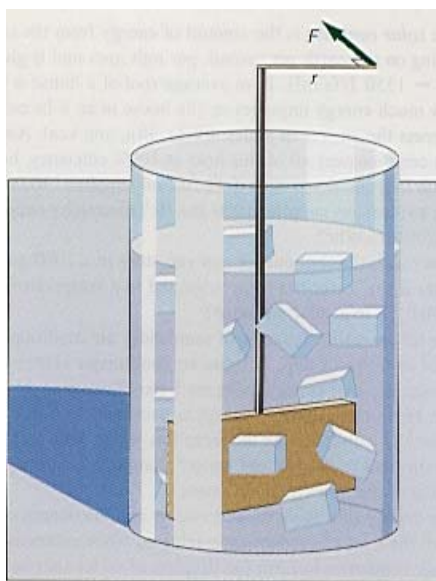


Diagram for problem 40.

*41. A 75.0-kg patient is running a fever of 105 °F (40.6 °C) and is given an alcohol rub down to lower the body temperature. If the specific heat of the human body is approximately 3474 J/(kg °C), and the heat of vaporization of alcohol is 8.50×10^5 J/kg, find (a) the amount of heat that must be removed to lower the temperature to 102 °F (38.9 °C) and (b) the volume of alcohol required.

42. How much thermal energy is required to heat the air in a house from 15.0 °C to 20.0 °C if the house is 14.0 m long, 9.00 m wide, and 3.00 m high?

43. A classroom is at an initial temperature of 20 °C. If 35 students enter the class and each liberates heat to the air at the rate of 100 W, find the final temperature of the air in the room 50 min later, assuming all the heat from the students goes into heating the air. The classroom is 10.0 m long, 9.00 m wide, and 4.00 m high.

44. How much fuel oil is needed to heat a 570-liter tank of water from 10.0 °C to 80.0 °C if oil is capable of supplying 3.88×10^7 J of thermal energy per liter of oil?

45. How much heat is necessary to melt 100 kg of aluminum initially at a temperature of 20 °C? The melting point of aluminum is 660 °C and its heat of fusion is 3.77×10^5 J/kg.

*46. If the heat of combustion of natural gas is 3.71×10^7 J/m³, how many cubic meters are needed to heat 0.580 m³ of water from 10.0 °C to 75.0 °C in a hot water heater if the system is 63% efficient?

*47. If the heat of combustion of coal is 2.78×10^7 J/kg, how many kilograms of coal are necessary to heat 0.580 m³ of water from 10.0 °C to 75.0 °C in a hot water heater if the system is 63% efficient?

*48. The *solar constant* is the amount of energy from the sun falling on the earth per second, per unit area and is given as $S_c = 1350$ J/(s m²). If an average roof of a

house is 60.0 m², how much energy impinges on the house in an 8-hr period? Express the answer in joules, kWhr, Btu, and kcal. Assuming you could convert all of this heat at 100% efficiency, how much fuel could you save if #2 fuel oil supplies 1.47×10^8 J/gal; natural gas supplies 3.71×10^7 J/m³; electricity supplies 3.59×10^6 J/kWhr?

49. How much thermal energy can you store in a 5680-liter tank of water if the water has been subjected to a temperature change of 35.0 °C in a solar collector?

50. A 5.94-kg lead ball rolls without slipping down a rough inclined plane 1.32 m long that makes an angle of 40.0° with the horizontal. The ball has an initial velocity $v_0 = 0.25$ m/s. The ball is not perfectly spherical and some energy is lost due to friction as it rolls down the plane. The ball arrives at the bottom of the plane with a velocity $v = 3.00$ m/s, and 80.0% of the energy lost shows up as a rise in the temperature of the ball. Find (a) the height of the incline, (b) the initial potential energy of the ball, (c) the initial kinetic energy of translation, (d) the initial kinetic energy of rotation, (e) the initial total energy of the ball, (f) the final kinetic energy of translation, (g) the final kinetic energy of rotation, (h) the final total mechanical energy of the ball at the bottom of the plane, (i) the energy lost by the ball due to friction, and (j) the increase in the temperature of the ball.

*51. The energy that fuels thunderstorms and hurricanes comes from the heat of condensation released when saturated water vapor condenses to form the droplets of water that become the clouds that we see in the sky. Consider the amount of air contained in an imaginary box 5.00 km long, 10.0 km wide, and 30.0 m high that covers the ground at the surface of the earth at a particular time. The air temperature is 20 °C

and is saturated with all the water vapor it can contain at that temperature, which is 17.3×10^{-3} kg of water vapor per m^3 . The air in this imaginary box is now lifted into the atmosphere where it is cooled to 0°C . Since the air is saturated, condensation occurs throughout the cooling process. The maximum water vapor the air can contain at 0°C is 4.847×10^{-3} kg of water vapor per m^3 . (The heat of vaporization of water varies with temperatures from 2.51×10^6 J/kg at 0°C to 2.26×10^6 J/kg at 100°C . We will assume an average temperature of 10.0°C for the cooling process.) Find (a) the volume of saturated air in the imaginary box, (b) the mass of water vapor in this volume at 20.0°C , (c) the mass of water vapor in this volume at 0°C , (d) the heat of vaporization of water at 10.0°C , and (e) the thermal energy given off in the condensation process. (f) Discuss this quantity of energy in terms of the energy that powers thunderstorms and hurricanes.

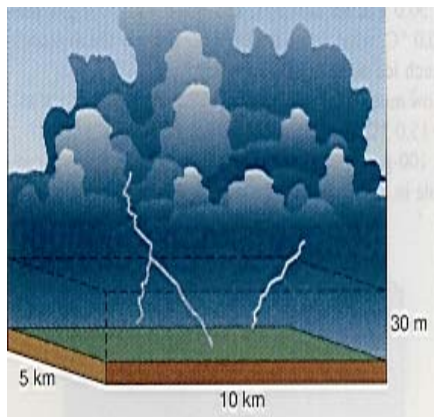


Diagram for problem 51.

Interactive Tutorials

52. *Convert ice to water.* Find the total amount of thermal energy in joules necessary to convert ice of mass $m_i = 2.00$ kg at an initial temperature $t_{ii} = -20.0^\circ\text{C}$ to water at a final water temperature of $t_{fw} = 88.3^\circ\text{C}$. The specific heat of ice is $c_i = 2093$ J/kg $^\circ\text{C}$, water is $c_w = 4186$ J/kg $^\circ\text{C}$, and the latent heat of fusion of water is $L_f = 3.34 \times 10^5$ J/kg.

53. *Equilibrium.* If a sample of lead shot of mass $m_s = 0.080$ kg and initial temperature $t_{is} = 100^\circ\text{C}$ is placed into a mass of water $m_w = 0.100$ kg in an aluminum calorimeter of mass $m_c = 0.060$ kg at an initial temperature $t_{iw} = 20.0^\circ\text{C}$, what is the final equilibrium temperature of the water, calorimeter, and lead shot? The specific heats are water $c_w = 4186$ J/kg $^\circ\text{C}$, calorimeter $c_c = 900$ J/kg $^\circ\text{C}$, and lead sample $c_s = 129.8$ J/kg $^\circ\text{C}$.

54. *Temperature Conversion Calculator.* The Temperature Conversion Calculator will permit you to convert temperatures in one unit to a temperature in another unit.

55. *Specific heat.* A specimen of lead, $m_s = 0.250$ kg, is placed into an oven where it acquires an initial temperature $t_{is} = 200^\circ\text{C}$. It is then removed and placed into a calorimeter of mass $m_c = 0.060$ kg and specific heat $c_c = 900$ J/kg $^\circ\text{C}$ that contains water, $m_w = 0.200$ kg, at an initial temperature $t_{iw} = 10.0^\circ\text{C}$. The specific heat of water is $c_w = 4186$ J/kg $^\circ\text{C}$. The final equilibrium temperature of the water in the calorimeter is observed

to be $t_{fw} = 16.7^\circ\text{C}$. Find the specific heat c_s of this sample.

56. *Converting ice to superheated steam.* Find the total amount of thermal energy in joules necessary to convert ice of mass $m_i = 12.5$ kg at an initial temperature $t_{ii} = -25.0^\circ\text{C}$ to superheated steam at a temperature $t_{ss} = 125^\circ\text{C}$. The specific heat of ice is $c_i = 2093$ J/kg $^\circ\text{C}$, water is $c_w = 4186$ J/kg $^\circ\text{C}$, and steam is $c_s = 2013$ J/kg $^\circ\text{C}$. The latent heat of fusion of water is $L_f = 3.34 \times 10^5$ J/kg, and the latent heat of vaporization is $L_v = 2.26 \times 10^6$ J/kg.

57. *A mixture.* How much ice at an initial temperature of $t_{ii} = -15.0^\circ\text{C}$ must be added to a mixture of three specimens contained in a calorimeter in order to make the final equilibrium temperature of the water $t_{fw} = 12.5^\circ\text{C}$? The three specimens and their characteristics are sample 1: zinc; $m_{s1} = 0.350$ kg, $c_{s1} = 389$ J/kg $^\circ\text{C}$, initial temperature $t_{is1} = 150^\circ\text{C}$; sample 2: copper; $m_{s2} = 0.180$ kg, $c_{s2} = 385$ J/kg $^\circ\text{C}$, initial temperature $t_{is2} = 100^\circ\text{C}$; and sample 3: tin; $m_{s3} = 0.350$ kg, $c_{s3} = 226$ J/kg $^\circ\text{C}$, initial temperature $t_{is3} = 180^\circ\text{C}$. The calorimeter has a mass $m_c = 0.060$ kg and specific heat $c_c = 900$ J/kg $^\circ\text{C}$ and contains water, $m_w = 0.200$ kg, at an initial temperature $t_{iw} = 19.5^\circ\text{C}$. The specific heat of water is $c_w = 4186$ J/kg $^\circ\text{C}$.

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