TEMPERATURE, HEAT, AND EXPANSION

Objectives
- Define temperature in terms of molecular motion. (21.1)
- Describe how heat flows. (21.2)
- Describe how a thermometer works. (21.3)
- Explain the connection between internal energy and heat. (21.4)
- Describe how the quantity of heat that enters or leaves a substance is determined. (21.5)
- Explain why the specific heat capacities of different substances are different. (21.6)
- Describe how water’s high specific heat capacity affects climate. (21.7)
- Explain how matter changes when heated or cooled. (21.8)
- Explain why ice floats on water. (21.9)

THE BIG IDEA
When matter gets warmer, the atoms or molecules in the matter move faster.

All matter—solid, liquid, and gas—is composed of continually jiggling atoms or molecules. Because of this random motion, the atoms and molecules in matter have kinetic energy. The average kinetic energy of these individual particles causes an effect we can sense—warmth. Whenever something becomes warmer, the kinetic energy of its atoms or molecules has increased.

It’s easy to increase the kinetic energy in matter. You can warm a penny by striking it with a hammer—the blow causes the molecules in the penny to jostle faster. If you put a flame to a liquid, the liquid also becomes warmer. Rapidly compress air in a tire pump and the air becomes warmer. When the atoms or molecules in matter move faster, the matter gets warmer. Its atoms or molecules have more kinetic energy. For brevity in this chapter, rather than saying atoms and molecules, we’ll simply say molecules—by which we mean either.

So when you warm up by a fire on a cold winter night, you are increasing the molecular kinetic energy in your body.

discover!

MATERIALS balloons, water, matches
EXPECTED OUTCOME The air-filled balloon will expand; the water-filled balloon will not.
ANALYZE AND CONCLUDE
1. Nothing happened to the water-filled balloon; the air-filled balloon expanded
2. A long time!
3. Water has a large specific heat capacity and can absorb a great deal of heat with little change in temperature. Thus, the temperature at the surface of the water-filled balloon does not increase sufficiently to rupture the balloon.

How Much Heat Can a Balloon Hold?
1. Fill a balloon with air and fill a second, similar balloon with water.
2. Hold a lighted match close to the bottom of the air-filled balloon. CAUTION! Do not touch the match to the balloon. Remove the match if the balloon looks as if it is about to rupture.
3. Now hold a lighted match near the bottom of the water-filled balloon.

Analyze and Conclude
1. Observing What happened to each of the balloons when exposed to the flame?
2. Predicting How long do you think it would take for the water-filled balloon to rupture if the flame were not removed?
3. Making Generalizations What role does water play in preventing the rupture of the balloon?
21.1 Temperature

The quantity that tells how hot or cold something is compared with a standard is temperature. We express temperature by a number that corresponds to a degree mark on some chosen scale.

Nearly all matter expands when its temperature increases and contracts when its temperature decreases. A common thermometer measures temperature by showing the expansion and contraction of a liquid—usually mercury or colored alcohol—in a glass tube using a scale. Temperature is generally measured on one of three different scales.

**Celsius Scale** On the most widely used temperature scale, the Celsius scale, the number 0 is assigned to the temperature at which water freezes, and the number 100 to the temperature at which water boils (at standard atmospheric pressure). The gap between freezing and boiling is divided into 100 equal parts, called degrees.

**Fahrenheit Scale** On the temperature scale used commonly in the United States, the Fahrenheit scale, the number 32 designates the temperature at which water freezes, and the number 212 is assigned to the temperature at which water boils (at 1 atm). The Fahrenheit scale will become obsolete if and when the United States goes metric.

**Kelvin Scale** The scale used in scientific research is the SI scale—the Kelvin scale. Its degrees are the same size as the Celsius degree and are called “kelvins.” On the Kelvin scale, the number 0 is assigned to the lowest possible temperature—absolute zero. At absolute zero, a substance has no kinetic energy to give up. Zero on the Kelvin scale, or absolute zero, corresponds to −273°C on the Celsius scale. We will learn more about the Kelvin scale in Chapter 24.

**Scale Conversion** Arithmetic formulas can be used for converting from one temperature scale to another and are often popular in classroom exams. Such arithmetic exercises are not really physics, so we will not be concerned with them here. Besides, a conversion from Celsius to Fahrenheit, or vice versa, can be very closely approximated by simply reading the corresponding temperature from the side-by-side scales in Figure 21.1.

![FIGURE 21.1](image)

This thermometer measures temperature on both Fahrenheit and Celsius scales.

**Key Terms**
- temperature
- Celsius scale
- Fahrenheit scale
- Kelvin scale
- absolute zero

**Teaching Tip** Review Section 9.5, Kinetic Energy, before doing this section. This will prepare students for the idea that temperature is related to the KE of the molecules of a substance. Strictly speaking, temperature is directly proportional to the KE per molecule only in the case of ideal gases. However, here we take the view that temperature is related to molecular translational KE in most common substances.

**Teaching Tip** Describe how the increased jostling of molecules in a substance results in expansion, the basis for the common thermometer. Sketch an uncalibrated thermometer on the board, with its mercury bulb at the bottom. Describe how energy is transferred from the outer environment to the mercury. When a thermometer is placed in boiling water, the jostling of the water molecules transfers to the mercury, which then expands up in the tube. Explain that the calibrations on such a tube are purely arbitrary. Then, introduce the three common temperature scales.

**Teaching Tip** For students who are not familiar with the Celsius scale, point out that room temperature is about 20°C, body temperature is 37°C, and desert temperatures often exceed 40°C.

**Teaching Tip** Emphasize the absolute zero of temperature, 0 K—the temperature at which there is no available KE.
Ask Which of the two temperature scales is the more precise when temperatures are expressed to the nearest whole number? Degree marks are closer together on the Fahrenheit scale because there are 180 Fahrenheit degrees between the freezing and boiling points of water, and only 100 on the Celsius scale. The smaller degrees allow more precise readings when temperatures are expressed to the nearest whole number.

Teaching Tip Explain that the molecules in a cup of hot coffee move faster than those in a cup of cold coffee. Discuss the quantity of this energetic jostling and haphazard bumbling of molecules—temperature. Define temperature as the average KE of molecules or atoms in a substance.

Desert Ants
The surface temperatures of some deserts in Africa and central Asia reach 60°C (140°F). This is hot, but not too hot for a species of ant (Cataglyphis) that thrives at this searing temperature. These desert ants can forage for food at temperatures too high for lizards who eat them. Resistant to heat, these ants can withstand higher temperatures than any other creatures in the desert. They scavenge the desert surface for corpses of those who did not find cover in time, touching the hot sand as little as possible while often sprinting on four legs with two high in the air. Although their foraging paths zigzag over the desert floor, their return paths are almost straight lines to their nest holes. They attain speeds of 100 body lengths per second. During an average six-day life, most of these ants retrieve 15 to 20 times their weight in food.

Temperature and Kinetic Energy
Temperature is related to the random motions of the molecules in a substance. In the simplest case of an ideal gas, temperature is proportional to the average kinetic energy of molecular translational motion (that is, motion along a straight or curved path). In solids and liquids, where molecules are more constrained and have potential energy, temperature is more complicated. But it is still true that temperature is closely related to the average kinetic energy of translational motion of molecules.

The higher the temperature of a substance, the faster is the motion of its molecules. So the warmth you feel when you touch a hot surface is the kinetic energy transferred by molecules in the surface to molecules in your fingers.

Note that temperature is not a measure of the total kinetic energy of all the molecules in a substance. There is twice as much kinetic energy in 2 liters of boiling water as in 1 liter. But the temperatures of both liters of water are the same because the average kinetic energy of molecules in each is the same. Figure 21.2 shows that a bucket of warm water can contain more molecular kinetic energy than a cup of hot water.

CONCEPT CHECK: What is the relationship between the temperature of a substance and the speed of its molecules?
21.2 Heat

If you touch a hot stove, energy enters your hand from the stove because the stove is warmer than your hand. But if you touch ice, energy passes from your hand into the colder ice. The direction of spontaneous energy transfer is always from a warmer to a cooler substance. The energy that transfers from one object to another because of a temperature difference between them is called heat.

It is common—but incorrect with physics types—to think that matter contains heat. Matter contains energy in several forms, but it does not contain heat. Heat is energy in transit, moving from a body of higher temperature to one of lower temperature. Once transferred, the energy ceases to be heat.

In Chapter 9, we called the energy resulting from heat flow thermal energy, to make clear its link to heat and temperature. In this and following chapters, we will use the term that scientists prefer, internal energy.

When heat flows from one object or substance to another one it is in contact with, the objects or substances are in thermal contact. Figure 21.3 uses an analogy to show how heat flows between two objects in thermal contact. When two substances of different temperatures are in thermal contact, heat flows from the higher-temperature substance into the lower-temperature substance. However, heat will not necessarily flow from a substance with more total molecular kinetic energy to a substance with less. For example, there is more total molecular kinetic energy in a large bowl of warm water than there is in a red-hot thumbtack. Yet, if the tack is immersed in the water, heat does not flow from the water to the tack. It flows from the hot tack to the cooler water. Heat flows according to temperature differences—that is, average molecular kinetic energy differences. Heat never flows on its own from a cooler substance into a hotter substance.

CONCEPT CHECK What causes heat to flow?

discover!

Can You Trust Your Senses?

1. Put some hot water, some warm water, and some cold water in three open containers.
2. Place a finger in the hot water and a finger of the other hand in the cold water. How do they feel?
3. After a few seconds, place both fingers in the warm water. How do they feel now?
4. Think Why is a thermometer better for measuring temperature?
After objects in thermal contact with each other reach the same temperature, we say the objects are in **thermal equilibrium**. When objects are in thermal equilibrium, no heat flows between them.

To read a thermometer we wait until it reaches thermal equilibrium with the substance being measured. When a thermometer is in contact with a substance, heat flows between them until they have the same temperature. The temperature of the thermometer is also the temperature of the substance. So a thermometer, interestingly enough, shows only its own temperature. This is shown in Figure 21.4.

![FIGURE 21.4](somewhat-like-water-in-the-pipes-seeking-a-common-level-for-which-the-preserves-at-equal-elevations-are-the-same-the-thermometer-and-its-immediate-surroundings-achieve-a-common-temperature-at-which-the-average-kinetic-energy-per-particle-is-the-same-for-both.)

A thermometer should be small enough that it does not appreciably alter the temperature of the substance being measured. If you are measuring the temperature of room air, then the heat absorbed by the thermometer will not lower the air temperature noticeably. But if you are trying to measure the temperature of a drop of water, the temperature of the drop after thermal contact may be quite different from its initial temperature.

**CONCEPT CHECK**: How does a thermometer measure temperature?

**think!**

Suppose you use a flame to add a certain quantity of heat to 1 liter of water, and the water temperature rises by 2°C. If you add the same quantity of heat to 2 liters of water, by how much will its temperature rise? **Answer: 21.5.1**
21.4 Internal Energy

In addition to the translational kinetic energy of jostling molecules in a substance, there is energy in other forms. There is rotational kinetic energy of molecules and kinetic energy due to internal movements of atoms within molecules. There is also potential energy due to the forces between molecules. The grand total of all energies inside a substance is called **internal energy**. A substance does not contain heat—it has internal energy.

**Concept Check**: What happens to the internal energy of a substance that takes in or gives off heat?

21.5 Measurement of Heat

So we see that heat is energy transferred from one substance to another by a temperature difference. **The amount of heat transferred can be determined by measuring the temperature change of a known mass of a substance that absorbs the heat.**

When a substance absorbs heat, the resulting temperature change depends on more than just the mass of the substance, as shown in Figure 21.5. The quantity of heat that brings a cupful of soup to a boil might raise the temperature of a pot of soup by only a few degrees. To quantify heat, we must specify the mass and kind of substance affected.

The unit of heat is defined as the heat necessary to produce some standard, agreed-on temperature change for a specified mass of material. The most commonly used unit for heat is the **calorie**. The **calorie** is defined as the amount of heat required to raise the temperature of 1 gram of water by 1°C.
The calorie is a convenient heat unit because it makes the specific heat capacity of water numerically equal to 1. It is commonly used in chemistry. Your students’ future science courses may use the joule for the heat unit, in which case the calorie is a stepping stone to SI. You can teach in calories or joules, or both!

**think!**

Which will raise the temperature more, adding 1 calorie or 4.186 joules?

*Answer: 21.5.2*

**Demonstration**

Demonstrate the energy in food by spearing a walnut meat with a paper clip. Holding the paper clip with tongs, ignite the walnut under a beaker of 50 mL of water. The water will quickly boil.

**Ask** 1 L of 30°C water is mixed with 1 L of 20°C water. What is the final temperature of the mixture? 25°C; the internal energy lost by the warmer water is gained by the cooler water.

**Ask** Why do experts say not to refreeze food that has been thawed? If the food spends too much time above refrigeration temperature (40°F) it may be unsafe to eat due to bacteria growth. So you’d be refreezing unsafe food. If the food still contains ice crystals and is as cold as if it were refrigerated, it may be refrozen safely. But ice cream and frozen yogurt are exceptions and should be discarded. For food such as bread that never needed freezing in the first place, refreezing is safe.

**CONCEPT CHECK:** The amount of heat transferred can be determined by measuring the temperature change of a known mass of a substance that absorbs the heat.

![FIGURE 21.6 ▲](image)

To the weight watcher, the peanut contains 10 Calories; to the physicist, it releases 10,000 calories (or 41,860 joules) of energy when burned or digested.

The kilocalorie is 1000 calories (the heat required to raise the temperature of 1 kilogram of water by 1°C). The heat unit used in rating foods is actually a kilocalorie, although it’s often referred to as the calorie. To distinguish it from the smaller calorie, the food unit is sometimes called a Calorie (written with a capital C).

It is important to remember that the calorie and the Calorie are units of energy. In the International System of Units (SI), quantity of heat is measured in joules, the SI unit for all forms of energy. The relationship between calories and joules is that 1 calorie equals 4.186 J.

In this book, we’ll learn about heat with the conceptually simpler calorie—but in the lab, you may use the joule equivalent, where an input of 4.186 joules raises the temperature of 1 gram of water by 1°C.21.5

The energy value in food is determined by burning the food and measuring the energy that is released as heat. Food and other fuels are rated by how much energy a certain mass of the fuel gives off as heat when burned.

**CONCEPT CHECK:** How can you determine the amount of heat transferred to a substance?

**do the math!**

A woman with an average diet consumes and expends about 2000 Calories per day. The energy used by her body is eventually given off as heat. How many joules per second does her body give off?

We find this by converting 2000 Calories per day to joules per second. We use the information that 1 Calorie equals 4186 joules, 1 day equals 24 hours, and 1 hour equals 3600 seconds. The conversion is then set up as follows:

\[
\frac{2000 \text{ Cal}}{1 \text{ d}} \times \frac{1 \text{ d}}{24 \text{ hr}} \times \frac{1 \text{ hr}}{3600 \text{ s}} \times \frac{4186 \text{ J}}{1 \text{ Cal}} = 97 \text{ J/s} = 97 \text{ W}
\]

Notice that the original quantity (2000 Cal/d) is multiplied by a set of fractions in which the numerator equals the denominator. Since each fraction has the value 1, multiplying by it does not change the value of the original quantity. The rule for choosing which quantity to put in the numerator is that the units should cancel and reduce to those of the end result. (We call this technique “dimensional analysis.”) On the average, the woman emits heat at the rate of 97 J/s, which is 97 watts. This is nearly the same as a glowing 100-W lamp! It’s easy to see why a crowded room soon becomes warm!
21.6 Specific Heat Capacity

Almost everyone has noticed that some foods remain hot much longer than others. Boiled onions and moist squash on a hot dish, for example, are often too hot to eat while mashed potatoes may be just right. The filling of hot apple pie can burn your tongue while the crust will not, even when the pie has just been taken out of the oven. The aluminum covering on a frozen dinner can be peeled off with your bare fingers as soon as it is removed from the oven, as shown in Figure 21.7. (But be careful of the food beneath it!)

Different substances have different capacities for storing internal energy, or heat. The capacity of a substance to store heat depends on its chemical composition. If we heat a pot of water on a stove, we may find that it requires 15 minutes to raise it from room temperature to its boiling temperature. But if we were to put an equal mass of iron on the same flame, we would find that it would rise through the same temperature range in only about 2 minutes. For silver, the time would be less than a minute. A specific material requires a specific amount of heat to raise the temperature of a given mass of the material by a specified number of degrees. The specific heat capacity of a material is the quantity of heat required to raise the temperature of a unit mass of the material by 1 degree.

### Table 21.1 Specific Heat Capacities

<table>
<thead>
<tr>
<th>Material</th>
<th>(J/g°C)</th>
<th>(cal/g°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>4.186</td>
<td>1.00</td>
</tr>
<tr>
<td>Aluminum</td>
<td>0.900</td>
<td>0.215</td>
</tr>
<tr>
<td>Clay</td>
<td>1.4</td>
<td>0.33</td>
</tr>
<tr>
<td>Copper</td>
<td>0.386</td>
<td>0.092</td>
</tr>
<tr>
<td>Lead</td>
<td>0.128</td>
<td>0.031</td>
</tr>
<tr>
<td>Olive Oil</td>
<td>1.97</td>
<td>0.471</td>
</tr>
<tr>
<td>Silver</td>
<td>0.23</td>
<td>0.056</td>
</tr>
<tr>
<td>Steel (iron)</td>
<td>0.448</td>
<td>0.107</td>
</tr>
</tbody>
</table>

We can think of specific heat capacity as thermal inertia. Recall that inertia is a term used in mechanics to signify the resistance of an object to change in its state of motion. Specific heat capacity is like a thermal inertia since it signifies the resistance of a substance to change in its temperature.
Absorbed energy can affect substances in different ways. Absorbed energy that increases the translational speed of molecules is responsible for increases in temperature. Absorbed energy may also increase the rotation of molecules, increase the internal vibrations within molecules, or stretch intermolecular bonds and be stored as potential energy. These kinds of energy, however, are not measured by a substance’s temperature. Temperature is a measure only of the kinetic energy of translational motion. Generally, only part of the energy absorbed by a substance raises its temperature.

Whereas a gram of water requires 1 calorie of energy to raise the temperature 1°C, it takes only about one eighth as much energy to raise the temperature of a gram of iron by the same amount. Iron atoms in the iron lattice primarily shake back and forth in translational fashion, while water molecules soak up a lot of energy in rotations, internal vibrations, and bond stretching. So water absorbs more heat per gram than iron for the same change in temperature. Water has a higher specific heat capacity (sometimes simply called specific heat) than iron has.

**CONCEPT CHECK:** Why do different substances have different capacities to store heat?

**do the math!**

How many calories are needed to raise the temperature of 1 liter of water by 15°C?

When we know the specific heat capacity, \( c \), for a particular substance, then the quantity of heat, \( Q \), involved when the mass, \( m \), of the substance undergoes a temperature change, \( \Delta T \), is \( Q = mc\Delta T \). Heat transferred = mass \times specific heat capacity \times temperature change.

The specific heat capacity for water, \( c \), is 1 cal/g°C, and the mass of 1 liter of water is 1 kilogram, which is 1000 grams. Since \( c \) is expressed in calories per gram °C, we express the mass of water, \( m \), in grams.

Then,

\[
Q = mc\Delta T
\]

\[
Q = (1000 \text{ g})(1 \text{ cal/g°C})(15°) = 15,000 \text{ calories}
\]

Suppose we deliver this energy to the water with a 1000-watt immersion heater. How long will it take to heat the water?

We know that 1000 watts delivers energy at the rate 1000 joules per second. Converting calories to joules,

\[
15,000 \text{ cal} \times 4.186 \text{ J/cal} = 63,000 \text{ joules}
\]

At the rate of 1000 joules per second, the time required for heating the water by 15°C is 63 seconds, a little more than a minute.
21.7 The High Specific Heat Capacity of Water

Water has a much higher capacity for storing energy than most common materials. A relatively small amount of water absorbs a great deal of heat for a correspondingly small temperature rise. Because of this, water is a very useful cooling agent, and is used in cooling systems in automobiles and other engines. If a liquid of lower specific heat capacity were used in cooling systems, its temperature would rise higher for a comparable absorption of heat. (Of course, if the temperature of the liquid rises to the temperature of the engine, no further cooling will take place.) Water also takes longer to cool, a useful fact to your great-grandparents, who on cold winter nights likely used foot-warming hot-water bottles in their beds.

Water’s capacity to store heat also affects the global climate. As shown in Figure 21.8, water takes more energy to heat up than land does. ✔ The property of water to resist changes in temperature improves the climate in many places. Europe and the west coast of the United States both benefit from this property of water.

![Figure 21.8](image)
Water has a high specific heat and is transparent, so it takes more energy to heat up than land does.

**Climate of Europe**  The next time you are looking at a world globe, notice the high latitude of Europe. If water did not have a high heat capacity, the countries of Europe would be as cold as the northeastern regions of Canada, for both Europe and Canada get about the same amount of the sun’s energy per square kilometer. The Atlantic current known as the Gulf Stream brings warm water northeast from the Caribbean. It holds much of its internal energy long enough to reach the North Atlantic off the coast of Europe, where it then cools. The energy released (one calorie per degree for each gram of water that cools) is carried by the prevailing westerly winds over the European continent.
Climate of America  Similarly, the climates differ on the east and west coasts of North America. The prevailing winds in the latitudes of North America are westerly. On the west coast, air moves from the Pacific Ocean to the land. Because of water’s high heat capacity, ocean temperature does not vary much from summer to winter. The water is warmer than the air in the winter, and cooler than the air in the summer. In winter, the water warms the air that moves over it and warms the western coastal regions of North America. In summer, the water cools the air and the western coastal regions are cooled. On the east coast, air moves from the land to the Atlantic Ocean. Land, with a lower specific heat capacity, gets hot in summer but cools rapidly in winter. As a result of water’s high heat capacity and the wind directions, the west coast city of San Francisco is warmer in the winter and cooler in the summer than the east coast city of Washington, D.C., which is at about the same latitude.

The central interior of a large continent usually experiences extremes of temperature. For example, the high summer and low winter temperatures common in Manitoba and the Dakotas are largely due to the absence of large bodies of water. Europeans, islanders, and people living near ocean air currents should be glad that water has such a high specific heat capacity. San Franciscans are!

CONCEPT CHECK: What is the effect of water’s high specific heat capacity on climate?

21.8 Thermal Expansion

When the temperature of a substance is increased, its molecules jiggle faster and normally tend to move farther apart. This results in an expansion of the substance. Most forms of matter—solids, liquids, and gases—expand when they are heated and contract when they are cooled. You can see an example of this in Figure 21.9. For comparable pressures and comparable changes in temperature, gases generally expand or contract much more than liquids, and liquids expand or contract more than solids. This thermal expansion of solids must be accounted for in construction. It also has applications in certain electronic devices.

FIGURE 21.9

The extreme heat of a July day in Asbury Park, New Jersey, caused the buckling of these railroad tracks.
Expansion Joints If concrete sidewalks and highway paving were laid down in one continuous piece, cracks would appear due to the expansion and contraction brought about by the difference between summer and winter temperatures. To prevent this, the surface is laid in small sections, each one being separated from the next by a small gap that is filled in with a substance such as tar. On a hot summer day, expansion often squeezes this material out of the joints.

The expansion of materials must be allowed for in the construction of structures and devices of all kinds. Different materials expand at different rates. A dentist uses filling material that has the same rate of expansion as teeth. The aluminum pistons of an automobile engine are enough smaller in diameter than the steel cylinders to allow for the much greater expansion rate of aluminum. A civil engineer uses steel having the same expansion rate as concrete for reinforcing concrete. Long steel bridges often have one end fixed while the other rests on rockers that allow for expansion. The roadway itself is segmented with tongue-and-groove-type gaps called expansion joints, as shown in Figure 21.10.

Bimetallic Strips In a bimetallic strip, two strips of different metals, say one of brass and the other of iron, are welded or riveted together, as shown in Figure 21.11. When the strip is heated, the difference in the amounts of expansion of brass and iron shows up easily. One side of the double strip becomes longer than the other, causing the strip to bend into a curve. On the other hand, when the strip is cooled, it bends in the opposite direction, because the metal that expands the most also contracts the most. The movement of the strip may be used to turn a pointer, regulate a valve, or operate a switch.
Consider the expansion of a make-believe snugly fitting steel pipe that completely encircles Earth. How much longer would this 40-million-meter long pipe be if its temperature increased by 1°C?

Steel changes in length about 1 part in 100,000 for each Celsius degree change in temperature. This is a ratio, $\frac{1}{100,000}$.

For different lengths of steel, expansion follows the same proportion. For short lengths of steel, expansion may be negligible.

For the pipe, the ratio of its change in length $X$ to its full size is the same as the ratio above. For a 1°C temperature change:

$$\frac{1}{100,000} = \frac{X \text{ m}}{40,000,000 \text{ m}}$$

A little computation will show that the change in length $X$ is 400 m. Here's the interesting part: If such a pipe were elongated by this 400 m, then there would be a gap between it and Earth's surface. Would the gap be big enough to put this book under? To crawl under? To drive a truck under? How big would this gap be?

We can find the gap by ratio and proportion. The ratio of circumference $C$ to diameter $D$ for any circle is equal to $\pi$ (about 3.14). The ratio of the change in circumference $\Delta C$ to the change in diameter $\Delta D$ also has the same value:

$$\frac{\Delta C}{\Delta D} = \frac{400 \text{ m}}{\Delta D} = 3.14$$

Solve for $\Delta D$:

$$\Delta D = \frac{400 \text{ m}}{3.14} = 127.4 \text{ m}$$

This 127.4 m is the increase in diameter of the circular pipe. The size of the gap between Earth's surface and the expanded pipe is equal to the increase in radius, which is half the increase in diameter, or 63.7 m.

So if a steel pipe that fits snugly against Earth were increased in temperature by 1°C, perhaps by people all along its length breathing hard on it, the pipe would expand and stand an amazing 63.7 m off the ground! Using ratio and proportion is a straightforward way to solve many problems. Another way to solve for the expansion of a material involves a formula ($\Delta L = \alpha L_0 \Delta T$). You will encounter this formula in the lab part of your course.

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**Teaching Tip** Explain that different substances expand or contract (in length, area, and volume) at their own characteristic rates (coefficients of expansion). Explain the need for the same expansion rate in substances such as teeth and teeth fillings; iron reinforcing rods and concrete; and the metal wires that are encased in glass light bulbs and the glass itself.

**Ask** How would the calibration of a thermometer differ if the glass expanded more than the mercury? The scale would be upside down because the glass reservoir would expand and “open up”—like the enlarged hole in the heated metal ring—and allow more mercury to fill it. The mercury level would then fall with increases in temperature.

**Teaching Tip** Stress the importance of not filling a gasoline tank to the top: As gasoline expands in warm weather, overflow is a safety hazard (as well as a waste of money).
Thermostats  A thermostat, such as the one in Figure 21.12, is a practical application of a bimetallic strip that is used to control temperature. As the temperature of a room changes, the back-and-forth bending of the bimetallic coil opens and closes an electric circuit. When the room becomes too cold, the coil bends toward the brass side, and in so doing it closes an electric switch that turns on the heat. When the room becomes too warm, the coil bends toward the iron side, which opens the switch and turns off the heating unit. Refrigerators are equipped with special thermostats to prevent them from becoming too hot or too cold. Bimetallic strips are used in oven thermometers, electric toasters, and other devices.

Glass  How much a substance expands depends on its change in temperature. If one part of a piece of glass is heated or cooled more rapidly than adjacent parts, the expansion or contraction that results may break the glass. This is especially true for thick glass. Borosilicate glass is formulated to expand very little with increasing temperature.

CONCEPT CHECK  How does matter change when heated or cooled?

![How Can You Open a Tightly Closed Jar?](discover.png)

<table>
<thead>
<tr>
<th><strong>How Can You Open a Tightly Closed Jar?</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Find a glass jar with a metal lid that is difficult to open.</td>
</tr>
<tr>
<td>2. Heat the lid by placing it in a stream of hot water or momentarily placing it on a hot stove. Try to unscrew the lid. What happens?</td>
</tr>
<tr>
<td>3. <strong>Think</strong> Why is the jar easier to open after the metal lid is heated?</td>
</tr>
</tbody>
</table>

21.9 Expansion of Water

Almost all liquids will expand when they are heated. Ice-cold water, however, does just the opposite! Water at the temperature of melting ice, 0°C (or 32°F), contracts when the temperature is increased. This is most unusual. As the water is heated and its temperature rises, it continues to contract until it reaches a temperature of 4°C. With further increase in temperature, the water then begins to expand; the expansion continues all the way to the boiling point, 100°C. This odd behavior is shown graphically in Figure 21.13 on the next page.
A given amount of water has its smallest volume—and thus its greatest density—at 4°C. The same amount of water has its largest volume—and smallest density—in its solid form, ice. (Remember, ice floats in water, so it must be less dense than water.) The volume of ice at 0°C is not shown in Figure 21.13. (If it were plotted to the same exaggerated scale, the graph would extend far beyond the top of the page.) After water has turned to ice, further cooling causes it to contract.

The explanation for this behavior of water has to do with the odd crystal structure of ice. The crystals of most solids are structured so that the solid state occupies a smaller volume than the liquid state. Ice, however, has open-structured crystals, as shown in Figure 21.14. These crystals result from the angular shape of the water molecules, plus the fact that the forces binding water molecules together are strongest at certain angles. Water molecules in this open structure occupy a greater volume than they do in the liquid state. At 0°C, ice is less dense than water, and so ice floats on water.
Melting Ice  When ice melts, not all the open-structured crystals collapse. Some crystals remain in the ice-water mixture, which makes up a microscopic slush that slightly “bloats” the water (increasing its volume slightly). Ice water is therefore less dense than slightly warmer water. With an increase in temperature, more of the remaining ice crystals collapse. The melting of these crystals further decreases the volume of the water.

While crystals are collapsing as the temperature increases between 0°C and 10°C, increased molecular motion results in expansion. This effect is shown in the center graph in Figure 21.16. Whether ice crystals are in the water or not, increased vibrational motion of the molecules increases the volume of the water.

When we combine the effects of contraction and expansion, the curve looks like the right-hand graph in Figure 21.16 (or Figure 21.13). This behavior of water is of great importance in nature. Suppose that the greatest density of water were at its freezing point, as is true of most liquids. Then the coldest water would settle to the bottom, and ponds would freeze from the bottom up. Pond organisms would then be destroyed in winter months. Fortunately, this does not happen. The densest water, which settles at the bottom of a pond, is 4 degrees above the freezing temperature. Water at the freezing point, 0°C, is less dense and floats, so ice forms at the surface while the pond remains liquid below the ice.

Teaching Tip  Discuss the formation of ice, why it forms at the surface, and why it floats. Then explain that deep bodies of water don’t freeze over in winter because all the water in the lake has to be cooled to 4°C before colder water will remain at the surface to be cooled to the freezing temperature, 0°C. State that before one can cool a teaspoon full of the lake water to 3°C, let alone 0°C, all the water beneath the surface must be cooled to 4°C, and that winters are neither cold nor long enough for this to happen in the United States.
Freezing Water  Let’s examine this in more detail. Most of the cooling in a pond takes place at its surface, when the surface air is colder than the water. As the surface water is cooled, it becomes denser and sinks to the bottom. Water will “float” at the surface for further cooling only if it is as dense or less dense than the water below.

Consider a pond that is initially at, say, 10°C. It cannot possibly be cooled to 0°C without first being cooled to 4°C. And water at 4°C cannot remain at the surface for further cooling unless all the water below has at least an equal density— that is, unless all the water below is at 4°C. If the water below the surface is any temperature other than 4°C, any surface water at 4°C will be denser and sink before it can be further cooled. So before any ice can form, all the water in a pond must be cooled to 4°C. Only when this condition is met can the surface water be cooled to 3°, 2°, 1°, and 0°C without sinking. Then ice can form, as shown in Figure 21.17.

Thus, the water at the surface is first to freeze. Continued cooling of the pond results in the freezing of the water next to the ice, so a pond freezes from the surface downward. In a cold winter the ice will be thicker than in a milder winter.

Very deep bodies of water are not ice-covered even in the coldest of winters. This is because all the water in a lake must be cooled to 4°C before lower temperatures can be reached, and the winter is not long enough for all the water to be cooled to 4°C. If only some of the water is 4°C, it will lie on the bottom. Because of water’s high specific heat and poor ability to conduct heat, the bottom of deep lakes in cold regions is a constant 4°C the year round. Fish should be glad that this is so.

**Teaching Tip** Place an ice cube in a glass of ice water and compare the water level on the side of the glass before and after the ice melts. Ask your students to account for the volume of ice that extends above the water line before it melts. They should find that the level remains unchanged. The floating ice cube displaces its own weight of water, so if the cube weighs 1 N, then it displaces 1 N of water when placed in the glass, and the water level rises. If the cube is first melted and then poured into the glass, the water level would be higher, but by 1 N, the same amount. Account for the volume of floating ice that extends above the water line by explaining that the ice expanded upon freezing because of the hexagonal open structures of the crystals. The volume of all those billions and billions of open spaces equals the volume of ice extending above the water line! When the ice melts, the part above the water line fills in the open structures as the crystals collapse. Discuss this idea in terms of icebergs, and whether or not the coastline would change if all the floating icebergs in the world melted. The oceans would rise a bit, but only because icebergs are composed of fresh water. The slight rise is more easily understood by exaggerating the circumstance— think of ice cubes floating in mercury. When they melt, the depth of fluid (water on mercury) is higher than before.

**CONCEPT CHECK** At 0°C, ice is less dense than water, and so ice floats on water.
Concept Summary

- The higher the temperature of a substance, the faster is the motion of its molecules.
- When two substances are in thermal contact, heat flows from the higher-temperature substance into the lower-temperature one.
- When a thermometer is in contact with a substance, heat flows between them until they have the same temperature.
- When a substance takes in or gives off heat, its internal energy changes.
- Heat transferred can be found by measuring the temperature change of a known mass of substance that absorbs the heat.
- The capacity of a substance to store heat depends on its chemical composition.
- The property of water to resist changes in temperature improves the climate.
- Most forms of matter expand when they are heated and contract when they are cooled.
- At 0°C, ice is less dense than water, and so ice floats on water.

Key Terms

- temperature (p. 407)
- Celsius scale (p. 407)
- Fahrenheit scale (p. 407)
- Kelvin scale (p. 407)
- absolute zero (p. 407)
- heat (p. 409)
- thermal contact (p. 409)
- thermal equilibrium (p. 410)
- internal energy (p. 411)
- calorie (p. 411)
- kilocalorie (p. 412)
- specific heat capacity (p. 413)
- bimetallic strip (p. 417)
- thermostat (p. 419)

think! Answers

21.5.1 Its temperature will rise by 1°C because there are twice as many molecules in 2 liters of water and each molecule receives only half as much energy on average. So average kinetic energy, and temperature, increase by half as much.

21.5.2 Both are the same. This is like asking which is longer, a 1-mile-long track or a 1.6-kilometer-long track. They’re the same quantity expressed in different units.

21.6 Water has a greater heat capacity than sand. Water is much slower to warm in the hot sun and slower to cool in the cold night. Water has more thermal inertia. Sand’s low heat capacity, as evidenced by how quickly the surface warms in the morning sun and how quickly it cools at night, affects local climates.

21.8 Telephone lines are longer in summer, when they are warmer, and shorter in winter, when they are cooler. They therefore sag more on hot summer days than in winter. If the telephone lines are not strung with enough sag in summer, they might contract too much and snap during the winter.
Check Concepts

1. Temperature is closely related to average KE of molecular translational motion.
2. 100; 180
3. Temperature same because average KE in both the same
4. Matter contains internal energy. Heat is the flow of that energy due to a difference in temperature.
5. From high to low temperature
6. Its temperature reaches equilibrium with the temperature of the surroundings.
7. The state at which all temperatures are equal
8. The grand total of all the energy in a substance
9. 1,000 calories = 1 Calorie
10. 1 calorie = 4.184 joules. Both are units of energy.
11. It has a high or low ability to store internal energy.
12. Low
13. It is high.
14. The ocean gives energy to the air above it in the winter, which is then carried over land, warming it. In the summer, the water absorbs energy from the air above it, and the cooler air moves over the land and absorbs energy from it.
15. One side expands (and contracts) more than the other.
16. Gases
17. 4°C
18. Because of the presence of microscopic slush
19. Ice water and ice are less dense and float on the surface.

Section 21.1

1. What is the connection between temperature and kinetic energy?
2. How many degrees are between the melting point of ice and boiling point of water on the Celsius scale? Fahrenheit scale?
3. Why does 2 liters of boiling water not have twice as great a temperature as 1 liter of boiling water?

Section 21.2

4. Why is it incorrect to say that matter contains heat?
5. In terms of differences in temperature between objects in thermal contact, in what direction does heat flow?

Section 21.3

6. What is meant by saying that a thermometer measures its own temperature?
7. What is thermal equilibrium?

Section 21.4

8. What is internal energy?

Section 21.5

9. What is the difference between a calorie and a Calorie?

Section 21.6

10. What is the difference between a calorie and a joule?

Section 21.7

11. What does it mean to say that a material has a high or low specific heat capacity?
12. Do substances that heat up quickly normally have high or low specific heat capacities?

Section 21.8

13. How does the specific heat capacity of water compare with that of other common substances?
14. Why is the North American west coast warmer in winter months and cooler in summer months than the east coast?

Section 21.9

15. Why does a bimetallic strip curve when it is heated (or cooled)?
16. Which expands most for increases in temperature: solids, liquids, or gases?
17. At what temperature is the density of water greatest?
18. Ice is less dense than water because of its open crystalline structure. But why is water at 0°C less dense than water at 4°C?
19. Why do lakes and ponds freeze from the top down rather than from the bottom up?
20. Why do shallow lakes freeze quickly in winter, and deep lakes not at all?

**Think and Rank**

*Rank each of the following sets of scenarios in order of the quantity or property involved. List them from left to right. If scenarios have equal rankings, then separate them with an equal sign. (e.g., A = B)*

21. The four plastic-foam soup bowls contain the same amount of water at 20°C. You also have a batch of 100-g copper cylinders that have initial temperatures as shown. The cylinders are submerged in the bowls.

Rank the bowls according to the maximum temperature of the water after the cylinders are added.

22. Four plastic-foam soup bowls contain the same amount of water at 20°C. You dunk cylinders of different metals, but of equal masses, in the bowls. All four cylinders have been in a hot oven and have the same temperature. (See Table 21.1 for specific heat capacities of these metals.)

Rank the bowls according to the maximum temperature of the water after the cylinders are added.

**Plug and Chug**

Heat transfer in calories is given by \( Q = mc \Delta T \), where \( m \) is mass in grams, \( c \) is specific heat capacity in cal/g°C, and \( \Delta T \) is in °C.

23. Calculate the number of calories of heat needed to change 500 grams of water by 50 Celsius degrees.
24. Calculate the number of calories given off by 500 grams of water cooling from 50°C to 20°C.

25. A 30-gram piece of iron is heated to 100°C and then dropped into cool water where the iron's temperature drops to 30°C. How many calories does it lose to the water? (The specific heat capacity of iron is 0.11 cal/g°C.)

26. Suppose a 30-gram piece of iron is dropped into a container of water and gives off 165 calories in cooling. Calculate the iron's temperature change.

27. What mass of water will give up 240 calories when its temperature drops from 80°C to 68°C?

28. When a 50-gram piece of aluminum at 100°C is placed in water, it loses 735 calories of heat while cooling to 30°C. Calculate the specific heat capacity of the aluminum.

Think and Explain

29. a. None lower than air temperature
   b. Living things more
   c. Inanimate things same
   30. Celsius degrees are 9/5 as large.
   31. Yes, but the increase in water temperature would be negligible.
   32. No, a difference of 273 in 10,000,000 is insignificant.
   33. Iceberg, because more volume
   34. Pizza sauce (which contains a lot of water) has a higher specific heat capacity than the crust, and so the sauce gives off more heat per gram than the crust for an equal decrease in temperature.

24. \[ Q = mc\Delta T = (500 \text{ g}) \times (1 \text{ cal/g°C})(30°C) = 15,000 \text{ cal} \]
25. \[ Q = mc\Delta T = (30 \text{ g}) \times (0.11 \text{ cal/g°C})(70°C) = 231 \text{ cal} \]
26. \[ \Delta T = Q/mc = 165 \text{ cal/} [(30 \text{ g})(0.11 \text{ cal/g°C})] = 50°C \]
27. \[ m = Q/c\Delta T = 240 \text{ cal/} [(1.0 \text{ cal/g°C})(12°C)] = 20 \text{ g} \]
28. \[ c = Q/m\Delta T = 735 \text{ cal/} [(50 \text{ g})(70°C)] = 0.21 \text{ cal/g°C} \]
35. In the old days, on a cold winter night it was common to bring a hot object to bed with you. Which would be better—a 10-kilogram iron brick or a 10-kilogram jug of hot water at the same temperature? Explain.

36. In addition to the overall motion of molecules that is associated with temperature, some molecules can absorb large amounts of energy in the form of internal vibrations and rotations of the molecules themselves. Would you expect materials composed of such molecules to have a high or a low specific heat capacity? Why?

37. Desert sand is very hot in the day and very cool at night. What does this tell you about its specific heat?

38. Why does adding the same amount of heat to two different objects not necessarily produce the same increase in temperature?

39. When a 1-kg metal pan containing 1 kg of cold water is removed from the refrigerator and set on a table, which absorbs more heat from the room—the pan or the water? Defend your answer.

40. On a hot day, you remove from a picnic cooler a chilled watermelon and some chilled sandwiches. Which will remain cool for a longer time? Why?

41. Why is it important to protect water pipes so they don’t freeze?

42. Iceland, so named to discourage conquest by expanding empires, is not at all ice-covered like Greenland and parts of Siberia, even though it is nearly on the Arctic Circle. The average winter temperature of Iceland is considerably higher than regions at the same latitude in eastern Greenland and central Siberia. Why is this so?

43. A metal ball is just able to pass through a metal ring. When the ball is heated, thermal expansion will not allow it to pass through the ring. What would happen if the ring, rather than the ball, were heated? Would the ball pass through the heated ring? Does the size of the hole in the ring increase, decrease, or stay the same?

44. After a machinist slips a hot, snugly fitting iron ring over a cold brass cylinder, the ring becomes “locked” in position and can’t be removed even by subsequent heating. This procedure is called “shrink fitting.” How does it occur? Can you conclude anything about the thermal expansion rates of iron and brass?

35. Water has the higher specific heat capacity so it gives off more heat as it cools to room temperature.

36. High, because they have more ways in which to store energy.

37. Sand has a low specific heat.

38. Different specific heat capacities

39. The water absorbs more, due to higher specific heat. Same mass, so higher specific heat means more absorbed heat.

40. Watermelon, with its high water content and therefore high specific heat capacity, undergoes the least temperature change for a given gain of heat.

41. Water expands when it freezes and can rupture pipes.

42. Iceland is warmed by surrounding water.

43. The hole expands; yes

44. When the iron is heated, it expands—but the brass expands more, preventing the ring from coming loose.
45. Suppose you cut a small gap in a metal ring, as shown. If you heat the ring, will the gap become wider or narrower?

46. Would a bimetallic strip function if the two different metals happened to have the same rates of expansion? Is it important that they expand at different rates? Explain.

47. Cite an exception to the claim that all substances expand when heated.

48. An old remedy for a pair of nested drinking glasses that stick together is to run water at different temperatures into the inner glass and over the surface of the outer glass. Which water should be hot, and which cold?

49. State whether water at the following temperatures will expand or contract when warmed: 0°C; 4°C; 6°C.

50. Suppose that water is used in a thermometer instead of mercury. If the temperature is at 4°C and then changes, why can’t the thermometer indicate whether the temperature is rising or falling?

51. If water had a lower specific heat capacity, would lakes be more likely or less likely to freeze in the winter?

52. How does the combined volume of the billions and billions of hexagonal open spaces in the crystals in a piece of ice compare with the portion of ice that floats above the surface of the water?

**Think and Solve**

53. People in the pioneering days placed hot potatoes in their pockets on cold winter days to keep their hands warm. Assuming that a potato is mostly water, Andrew calculates 24,000 calories of heat are released by a 350-g potato that cools from 85°C to 15°C. Alexis calculates that 102,000 joules of heat are released. Whose answer do you agree with, and why?

54. If you wished to warm 100 kg of water by 15°C for your bath, how much heat would be required? (Give your answer in calories and joules.)

55. Anthony’s thin plastic water bottle holds 500 mL of water at temperature 28°C. He puts it into the refrigerator. Show that the refrigerator removes 50,000 J of heat from the water to cool it to 4°C.
56. Samantha decides to try the “Ice-Water Diet.” She drinks water at 0°C, and it must warm up to her body temperature, 37°C.
   a. How much ice water must she drink to “burn” 3500 Calories (the approximate energy content of one pound of fat)? Each Calorie is 1000 calories.
   b. Why would she recommend, or not recommend, this diet for losing weight?

57. What would be the final temperature of the mixture of 50 g of 20°C water and 50 g of 40°C water?

58. What would be the final temperature if you mixed a liter of 20°C water with 2 liters of 40°C water?

59. What would be the final temperature if you mixed a liter of 40°C water with 2 liters of 20°C water?

60. What would be the final temperature when 100 g of 25°C water is mixed with 75 g of 40°C water?

61. What will be the final temperature of 100 g of 20°C water when 100 g of 40°C iron nails are submerged in it? (The specific heat of iron is 0.12 cal/g°C.)

62. What is the specific heat capacity of a 50-gram piece of 100°C metal that will change 400 grams of 20°C water to 22°C?

63. Taylor finds that a certain amount of heat raises the temperature of a sample of iron by 10°C. Show that the same amount of heat will raise the temperature of an equal mass of lead by 35°C.

64. Suppose that a metal bar 1 m long expands 0.5 cm when it is heated. How much would it expand if it were 100 m long?

65. Steel expands 1 part in 100,000 for each Celsius degree increase in temperature. If the 1.5-km main span of a steel suspension bridge had no expansion joints, how much longer would it be for a temperature increase of 20°C?

66. A cook pours 1 L of ice water at 0°C into a pan of hot water at 80°C and finds that the mixture reaches a temperature of 60°C. How much hot water was in the pan?

67. Your (perfectly insulated) bathtub has 82 liters of water in it, but it has cooled down to a temperature of 39°C. You’d like to add just the right amount of 50°C water to the tub to make your bathwater the perfect temperature of 42°C. Show that adding 31 liters of 50°C water to the tub will accomplish this goal.

68. \[ Q_{\text{lost}} = Q_{\text{gained}}; mc \times (T - 20°C) = 2mc(40°C - T); \]
   \[ T = 33.3°C \]

69. \[ mc(40°C - T) = 2mc \times (T - 20°C); T = 26.7°C \]

70. \[ 100c(T - 25) = 75c \times (40 - T); T = 31.4°C \]

71. \[ 100c(T - 20) = 40(0.12) \times (40 - T); 100T - 2000 = 192 - 4.87; T = 20.9°C \]

72. \[ c = Q/m\Delta T = (400 \text{ g}) \times (1 \text{ cal/g°C})(22°C - 20°C)/[(50 \text{ g}) \times (100°C - 22°C)] = 0.2 \text{ cal/g°C} \]

73. Since \[ Q_{\text{iron}} = Q_{\text{lead}} \] and \[ m_{\text{iron}} = m_{\text{lead}} \], from \[ c_{\text{iron}}m_{\text{iron}}\Delta T_{\text{iron}} = c_{\text{lead}}m_{\text{lead}}\Delta T_{\text{lead}} \]; \[ \Delta T_{\text{lead}} = (c_{\text{iron}}/c_{\text{lead}}) \times \Delta T_{\text{iron}} = [(0.448 \text{ J/g°C})/(0.128 \text{ J/g°C})] \times 10°C = 35°C \]

74. \[ 0.5 \text{ cm}/(1 \text{ m}) = \Delta L/(100 \text{ m}), \text{ so } \Delta L = 50 \text{ cm} \]

75. \[ \Delta L = aL_{2}\Delta T = (10^{-5} \text{ m}) \times (1.5 \text{ km}) \times (20°C) = 3 \times 10^{-4} \text{ km} = 30 \text{ cm} \]

76. Heat lost by the hot water is gained by the ice water. Since the 20°C change of temperature of the hot water is only one-third as much as the change of the temperature of the cold water, there must be three times the mass of hot water. Therefore the pan contained 3 L of hot water.

77. From \[ Q_{\text{hot} - w} = Q_{\text{cold} - w} \] \[ m_{\text{hot} - w}\Delta T_{\text{hot} - w} = m_{\text{cold} - w}\Delta T_{\text{cold} - w} \]; So \[ m_{\text{hot} - w} = m_{\text{cold} - w}\Delta T_{\text{cold} - w} / \Delta T_{\text{hot} - w} = [82 \text{ kg} \times (42°C - 39°C)]/(50°C - 42°C) = 31 \text{ kg} = 31 \text{ L} \]