

PART
III

Properties of Matter



Individual bees do not have the ability to regulate their body temperature. But a colony of bees, working together, can very precisely regulate the temperature of their hive. How can the bees use the heat generated by their muscles, the structure of the hive, and the evaporation of water to achieve this control? In Part III, we will learn about the flows of matter and energy that drive such processes.

Beyond the Particle Model

The first 11 chapters of this book have made extensive use of the *particle model* in which we represent objects as point masses. The particle model is especially useful for describing how discrete objects move through space and how they interact with each other. Whether a ball is made of metal or wood is irrelevant to calculating its trajectory.

But there are many situations where the distinction between metal and wood is crucial. If you toss a metal ball and a wood ball into a pond, one sinks and the other floats. If you stir a pan on the stove with a metal spoon, it can quickly get too hot to hold unless it has a wooden handle.

Wood and metal have different physical properties. So do air and water. Our goal in Part III is to describe and understand the similarities and differences of different materials. To do so, we must go beyond the particle model and dig deeper into the nature of matter.

Macroscopic Physics

In Part III, we will be concerned with systems that are solids, liquids, or gases. Properties such as pressure, temperature, specific heat, and viscosity are characteristics of the system as a whole, not of the individual particles. Solids, liquids, and gases are often called *macroscopic* systems—the prefix “macro” (the opposite of “micro”) meaning “large.” We’ll make sense of the behavior of these macroscopic properties by considering a microscopic view in which we think of these systems as collections of particle-like atoms. This “micro-to-macro” development will be a key piece of the following chapters.

In the coming chapters, we’ll consider a wide range of practical questions, such as:

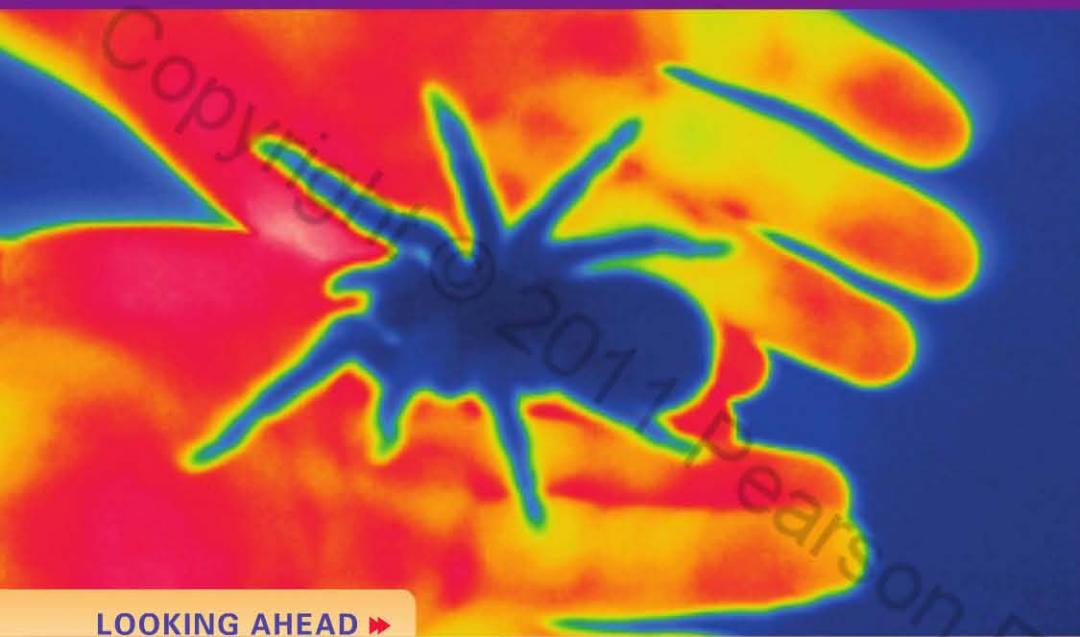
- How do the temperature and pressure of a system change when you heat it? Why do some materials respond quickly, others slowly?
- What are the mechanisms by which a system exchanges heat energy with its environment? Why does blowing on a cup of hot coffee cause it to cool off?
- Why are there three phases of matter—solids, liquids, gases? What happens during a phase change?
- Why do some objects float while others, with the same mass, sink? What keeps a massive steel ship afloat?
- What are the laws of motion of a flowing liquid? How do they differ from the laws governing the motion of a particle?

Both Newton’s laws and the law of conservation of energy will remain important tools—they are, after all, the basic laws of physics—but we’ll have to learn how they apply to macroscopic systems.

It should come as no surprise that an understanding of macroscopic systems and their properties is essential for understanding the world around us. Biological systems, from cells to ecosystems, are macroscopic systems exchanging energy with their environment. On a larger scale, energy transport on earth gives us weather, and the exchange of energy between the earth and space determines our climate.



12 Thermal Properties of Matter



LOOKING AHEAD ►

The goal of Chapter 12 is to use the atomic model of matter to explain and explore many macroscopic phenomena associated with heat, temperature, and the properties of matter.

The Ideal Gas

We'll continue to develop a model of the ideal gas based on the motion of individual atoms or molecules.

Pressure is due to collisions of molecules with the inside surface of the ball. More molecules or faster molecules mean more pressure.



Looking Back ◀

- 10.5 The nature of thermal energy
- 11.4 The ideal-gas model, the meaning of temperature

Ideal-Gas Processes

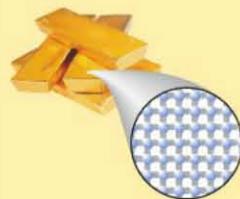
The pressure, volume, and temperature of a gas are related. Changing one means changing one or both of the others.



A Styrofoam cup is made of gas bubbles in plastic. The cup on the left was lowered deep into the ocean, where the huge pressure compressed the bubbles and deformed the cup.

The Atomic Model

The central theme that runs through this chapter is the atomic model. We will explain the thermal properties of matter by modeling solids, liquids, and gases in terms of atoms and the interactions between atoms.



We can explain a wide range of physical phenomena in terms of this simple underlying model.

Changes on Heating and Cooling

Heat causes both changes in temperature and changes of **phase**—solids change to liquid, liquids to gas.



Molten rock cools and solidifies when it hits the ocean. This transfer of heat causes the ocean water to boil.

Looking Back ◀

- 11.5 Heat and the first law of thermodynamics

Thermal Expansion

When liquids and solids are heated, they expand because of the increased motion of their atoms or molecules.



In Alaska, the heating of normally cool railroad tracks in the summer sun makes them expand and may create "sun kinks."

Heat Transfer

We'll explore the mechanisms by which heat is transferred from one object to another.



Air heated by the kettle streams upward, an example of heat transfer by convection.

12.1 The Atomic Model of Matter

We began exploring the concepts of thermal energy, temperature, and heat in Chapter 11, but many unanswered questions remain. How do the properties of matter depend on temperature? When you add heat to a system, by how much does its temperature change? And how is heat transferred to or from a system?

These are questions about the *macroscopic* state of systems, but we'll start our exploration by looking at a *microscopic* view, the atomic model that we've used to explain friction, elastic forces, and the nature of thermal energy. In this chapter, we'll use the atomic model to understand and explain the thermal properties of matter.

As you know, each element and most compounds can exist as a solid, liquid, or gas. These three **phases** of matter are familiar from everyday experience. An atomic view of the three phases is shown in **FIGURE 12.1**.

- A **gas** is a system in which each particle moves freely through space until, on occasion, it collides with another particle or the wall of its container.
- In a **liquid**, weak bonds permit motion while keeping the particles close together.
- A rigid **solid** has a definite shape and can be compressed or deformed only slightly, as we saw in Chapter 8. It consists of atoms connected by spring-like molecular bonds.

Our atomic model makes some simplifications that are worth noting. The basic particles of the gas in Figure 12.1 are drawn as simple spheres; no mention is made of the nature of the particles. The balloon might contain either helium (in which the basic particles are helium atoms) or air (in which the basic particles are nitrogen and oxygen molecules). A helium atom and a nitrogen molecule are quite different from each other, but many of the properties of the gas as a whole do not depend on the nature of the particles—a gas of helium atoms or oxygen molecules may behave identically. In such cases, we will simply refer to *gas particles*, which may be either atoms or molecules.

The basic particles in the liquid water in Figure 12.1 are water molecules. We ignore the structure of the molecules, so when we speak of the bonds that hold the particles in the liquid state, we are referring to the relatively weak intermolecular bonds among the water molecules, not the strong bonds between the hydrogen and oxygen atoms that form the molecules. The basic particles of the gold bars in Figure 12.1 are gold atoms. The bonds that hold these atoms together are the bonds that give the solid its structure. But there are solids composed of molecules that are held together by bonds between the molecules; water ice is one such example. In this case, the particles are the molecules, and the bonds that form the solid are the bonds between the molecules.

Atomic Mass and Atomic Mass Number

Before we see how the atomic model explains the thermal properties of matter, we need to remind you of some “atomic accounting.” Recall that atoms of different elements have different masses. The mass of an atom is determined primarily by its most massive constituents: the protons and neutrons in its nucleus. The *sum* of the number of protons and the number of neutrons is the **atomic mass number** *A*:

$$A = \text{number of protons} + \text{number of neutrons}$$

A, which by definition is an integer, is written as a leading superscript on the atomic symbol. For example, the primary isotope of carbon, with six protons (which makes it carbon) and six neutrons, has *A* = 12 and is written ^{12}C . The radioactive isotope ^{14}C , used for carbon dating of archeological finds, contains six protons and eight neutrons.

FIGURE 12.1 Atomic models of the three phases of matter: solid, liquid, and gas.

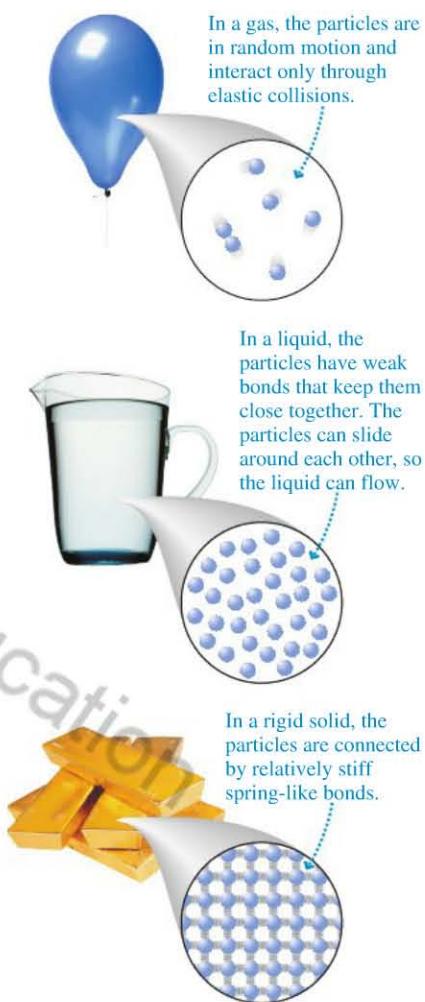


TABLE 12.1 Some atomic mass numbers

Element	Symbol	A
Hydrogen	^1H	1
Helium	^4He	4
Carbon	^{12}C	12
Nitrogen	^{14}N	14
Oxygen	^{16}O	16
Neon	^{20}Ne	20
Aluminum	^{27}Al	27
Argon	^{40}Ar	40
Lead	^{207}Pb	207

The **atomic mass** scale is established by defining the mass of ^{12}C to be exactly 12 u, where u is the symbol for the *atomic mass unit*. That is, $m(^{12}\text{C}) = 12 \text{ u}$. In kg, the atomic mass unit is

$$1 \text{ u} = 1.66 \times 10^{-27} \text{ kg}$$

Atomic masses are all very nearly equal to the integer atomic mass number A. For example, the mass of ^1H , with $A = 1$, is $m = 1.0078 \text{ u}$. For our present purposes, it will be sufficient to use the integer atomic mass numbers as the values of the atomic mass. That is, we'll use $m(^1\text{H}) = 1 \text{ u}$, $m(^4\text{He}) = 4 \text{ u}$, and $m(^{16}\text{O}) = 16 \text{ u}$. For molecules, the **molecular mass** is the sum of the atomic masses of the atoms that form the molecule. Thus the molecular mass of the diatomic molecule O_2 , the constituent of oxygen gas, is $m(\text{O}_2) = 2m(^{16}\text{O}) = 32 \text{ u}$.

NOTE ► An element's atomic mass number is *not* the same as its atomic number. The *atomic number*, which gives the element's position in the periodic table, is the number of protons. ◀

Table 12.1 lists the atomic mass numbers of some of the elements that we'll use for examples and homework problems. A complete periodic table, including atomic masses, is found in Appendix B.

The Definition of the Mole

One way to specify the amount of substance in a system is to give its mass. Another way, one connected to the number of atoms, is to measure the amount of substance in **moles**. **1 mole of substance, abbreviated 1 mol, is 6.02×10^{23} basic particles**.

The basic particle depends on the substance. Helium is a **monatomic gas**, meaning that the basic particle is the helium atom. Thus 6.02×10^{23} helium atoms are 1 mol of helium. But oxygen gas is a **diatomic gas** because the basic particle is the two-atom diatomic molecule O_2 . 1 mol of oxygen gas contains 6.02×10^{23} molecules of O_2 and thus $2 \times 6.02 \times 10^{23}$ oxygen atoms. Table 12.2 lists the monatomic and diatomic gases that we will use for examples and problems.

The number of basic particles per mole of substance is called **Avogadro's number** N_A . The value of Avogadro's number is thus

$$N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$$

The number n of moles in a substance containing N basic particles is

$$n = \frac{N}{N_A} \quad (12.1)$$

Moles of a substance in terms of the number of basic particles

The **molar mass** of a substance, M_{mol} , is the mass *in grams* of 1 mol of substance. To a good approximation, the numerical value of the molar mass equals the numerical value of the atomic or molecular mass. That is, the molar mass of He, with $m = 4 \text{ u}$, is $M_{\text{mol}}(\text{He}) = 4 \text{ g/mol}$, and the molar mass of diatomic O_2 is $M_{\text{mol}}(\text{O}_2) = 32 \text{ g/mol}$.

You can use the molar mass to determine the number of moles. In one of the few instances where the proper units are *grams* rather than kilograms, the number of moles contained in a system of mass M consisting of atoms or molecules with molar mass M_{mol} is

$$n = \frac{M \text{ (in grams)}}{M_{\text{mol}}} \quad (12.2)$$

Moles of a substance in terms of its mass



One mole of helium, sulfur, copper, and mercury.

EXAMPLE 12.1 Determining quantities of oxygen

A system contains 100 g of oxygen. How many moles does it contain? How many molecules?

SOLVE The diatomic oxygen molecule O₂ has molar mass $M_{\text{mol}} = 32 \text{ g/mol}$. From Equation 12.2,

$$n = \frac{100 \text{ g}}{32 \text{ g/mol}} = 3.1 \text{ mol}$$

Each mole contains N_A molecules, so the total number is $N = nN_A = 1.9 \times 10^{24}$ molecules.

Volume

An important property that characterizes a macroscopic system is its volume V , the amount of space the system occupies. The SI unit of volume is m³. Nonetheless, both cm³ and, to some extent, liters (L) are widely used metric units of volume. In most cases, you *must* convert these to m³ before doing calculations.

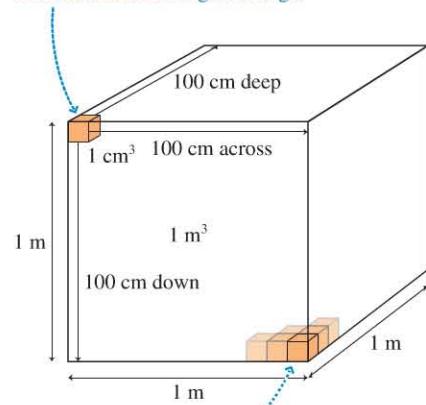
While it is true that 1 m = 100 cm, it is *not* true that 1 m³ = 100 cm³. FIGURE 12.2 shows that the volume conversion factor is 1 m³ = 10⁶ cm³. A liter is 1000 cm³, so 1 m³ = 10³ L. A milliliter (1 mL) is the same as 1 cm³.

STOP TO THINK 12.1 Which system contains more atoms: 5 mol of helium ($A = 4$) or 1 mol of neon ($A = 20$)?

- A. Helium. B. Neon. C. They have the same number of atoms.

FIGURE 12.2 There are 10⁶ cm³ in 1 m³.

Subdivide the 1 m × 1 m × 1 m cube into little cubes 1 cm on a side. You will get 100 subdivisions along each edge.



There are $100 \times 100 \times 100 = 10^6$ little 1 cm³ cubes in the big 1 m³ cube.

12.2 The Atomic Model of an Ideal Gas

Solids and liquids are nearly incompressible because the atomic particles are in close contact with each other. Gases, in contrast, are highly compressible because the atomic particles are far apart. In Chapter 11, we introduced an atomic-level model of an *ideal gas*, reviewed in FIGURE 12.3. Our goal in this section is to further develop this model to understand the pressure of an ideal gas.

Molecular Speeds and Temperature

The atomic model of an ideal gas is based on random motion, so it's no surprise that the individual atoms in a gas are moving at different speeds. FIGURE 12.4 on the next page shows data from an experiment to measure the molecular speeds in nitrogen gas at 20°C. The results are presented as a histogram, a bar chart in which the height of the bar indicates what percentage of the molecules have a speed in the range of speeds shown below the bar. For example, 16% of the molecules have speeds in the range from 600 m/s to 700 m/s. The most probable speed, as judged from the tallest bar, is ≈500 m/s. This is quite fast: ≈1200 mph!

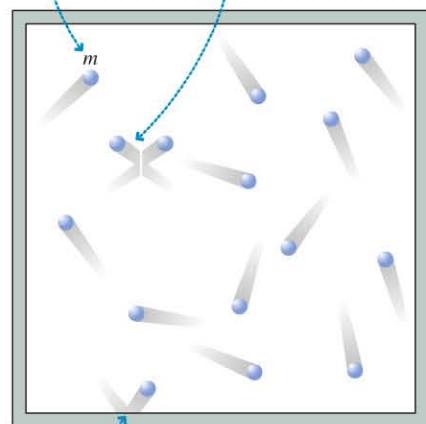
Because temperature is related to the average kinetic energy of the atoms, as we learned in Chapter 11, it will be useful to calculate the average kinetic energy for this distribution. An individual atom of mass m and velocity v has kinetic energy $K = \frac{1}{2}mv^2$. Recall that the average of a series of measurements is found by adding all the values and then dividing by the number of data points. Similarly, we can find the average kinetic energy by adding up all the kinetic energies of all the atoms and then dividing by the number of atoms:

$$K_{\text{avg}} = \frac{\sum \frac{1}{2}mv^2}{N} = \frac{1}{2}m \frac{\sum v^2}{N} = \frac{1}{2}m(v^2)_{\text{avg}} \quad (12.3)$$

FIGURE 12.3 The ideal-gas model.

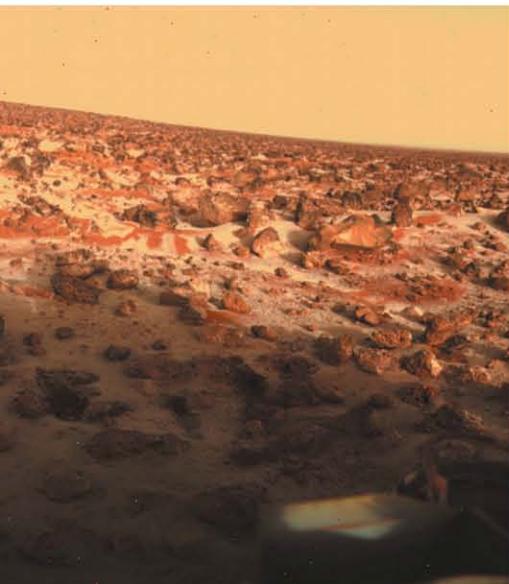
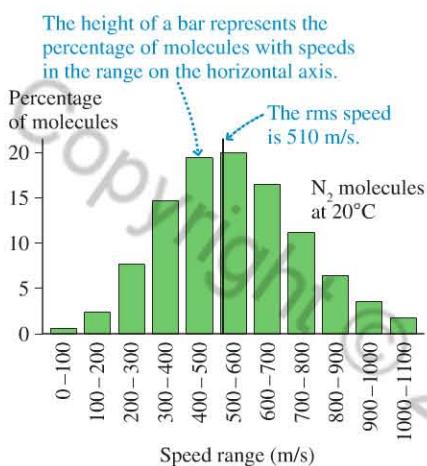
1. The gas is made up of a large number N of particles of mass m , each moving randomly.

2. The particles are quite far from each other and interact only rarely when they collide.



3. The collisions of the particles with each other (and with walls of the container) are elastic; no energy is lost in these collisions.

FIGURE 12.4 The distribution of molecular speeds in nitrogen gas at 20°C.



Martian airsicles The atmosphere of Mars is mostly carbon dioxide. At night, the temperature may drop so low that the molecules in the atmosphere will slow down enough to stick together—the atmosphere actually freezes. The frost on the surface in this image from the Viking 2 lander is composed partially of frozen carbon dioxide.

The quantity $\Sigma v^2/N$ is the sum of the values of v^2 for all the atoms divided by the number of atoms. By definition, this is average of the *squares* of all the individual speeds, which we've written $(v^2)_{\text{avg}}$.

The square root of this average is about how fast a typical atom in the gas is moving. Because we'll be taking the square root of the average, or mean, of the square of the speeds, we define the **root-mean-square speed** as

$$v_{\text{rms}} = \sqrt{(v^2)_{\text{avg}}} = \text{speed of a typical atom} \quad (12.4)$$

The root-mean-square speed is often referred to as the *rms speed*. The rms speed isn't the average speed of atoms in the gas; it's the speed of an atom with the average kinetic energy. But the average speed and the rms speed are very nearly equal, so we'll interpret an rms speed as telling us the speed of a typical atom in the gas.

Rewriting Equation 12.3 in terms of v_{rms} gives the average kinetic energy per atom:

$$K_{\text{avg}} = \frac{1}{2}mv_{\text{rms}}^2 \quad (12.5)$$

We learned in Chapter 11 that the temperature of a gas is related to the average kinetic energy of the atoms in the gas by

$$T = \frac{2}{3} \frac{K_{\text{avg}}}{k_B} \quad (12.6)$$

where $k_B = 1.38 \times 10^{-23}$ J/K is Boltzmann's constant. Indeed, Equation 12.4 is really the definition of temperature for an ideal gas—temperature measures the average kinetic energy of the particles in a system. We can now relate the temperature to the speeds of the atoms if we substitute Equation 12.5 into Equation 12.6:

$$T = \frac{1}{3} \frac{mv_{\text{rms}}^2}{k_B} \quad (12.7)$$

Solving Equation 12.7 for the rms speed of the atoms, we find that

$$v_{\text{rms}} = \sqrt{\frac{3k_B T}{m}} \quad (12.8)$$

NOTE ▶ You must use absolute temperature, in kelvin, to compute rms speeds. ◀

We have been considering ideal gases made of atoms, but our results are equally valid for real gases made of either atoms (such as helium, He) or molecules (such as oxygen, O₂). At a given temperature, Equation 12.8 shows that the speed of atoms or molecules in a gas varies with the atomic or molecular mass. A gas with lighter atoms will have faster atoms, on average, than a gas with heavier atoms. The equation also shows that higher temperatures correspond to faster atomic or molecular speeds. The rms speed is proportional to the *square root* of the temperature. This is a new mathematical form that we will see again, so we will take a look at its properties.

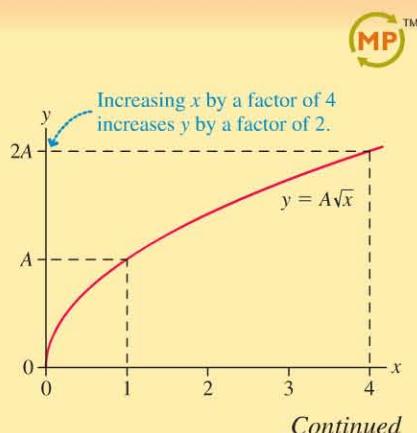
Square-root relationships

Two quantities are said to have a **square-root relationship** if y is proportional to the square root of x . We write the mathematical relationship as

$$y = A\sqrt{x}$$

y is proportional to the square root of x

The graph of a square-root relationship is a parabola that has been rotated by 90°.



Continued

SCALING If x has the initial value x_1 , then y has the initial value y_1 . Changing x from x_1 to x_2 changes y from y_1 to y_2 . The ratio of y_2 to y_1 is

$$\frac{y_2}{y_1} = \frac{A\sqrt{x_2}}{A\sqrt{x_1}} = \sqrt{\frac{x_2}{x_1}}$$

which is the square root of the ratio of x_2 to x_1 .

- If you increase x by a factor of 4, you increase y by a factor of $\sqrt{4} = 2$.
- If you decrease x by a factor of 9, you decrease y by a factor of $\sqrt{9} = 3$.

These examples illustrate a general rule:

Changing x by a factor of c changes y by a factor of \sqrt{c} .

Exercise 14

EXAMPLE 12.2 Speeds of air molecules

Most of the earth's atmosphere is the gas nitrogen, which consists of molecules, N_2 . At the coldest temperature ever observed on earth, $-129^\circ C$, what is the root-mean-square speed of the nitrogen molecules? Does the temperature at the earth's surface ever get high enough that a typical molecule is moving at twice this speed? (The highest temperature ever observed on earth was $58^\circ C$.)

PREPARE You can use the periodic table to determine that the mass of a nitrogen atom is 14 u. A molecule consists of two atoms, so its mass is 28 u. Thus the molecular mass in SI units (i.e., kg) is

$$m = 28 \text{ u} \times \frac{1.66 \times 10^{-27} \text{ kg}}{1 \text{ u}} = 4.6 \times 10^{-26} \text{ kg}$$

The problem statement gives two temperatures we'll call T_1 and T_2 ; we need to express these in kelvin. The lowest temperature ever observed on earth is $T_1 = -129 + 273 = 144 \text{ K}$; the highest temperature is $T_2 = 58 + 273 = 331 \text{ K}$.

SOLVE We use Equation 12.8 to find v_{rms} for the nitrogen molecules at T_1 :

$$v_{\text{rms}} = \sqrt{\frac{3k_B T_1}{m}} = \sqrt{\frac{3(1.38 \times 10^{-23} \text{ J/K})(144 \text{ K})}{4.6 \times 10^{-26} \text{ kg}}} = 360 \text{ m/s}$$

Because the rms speed is proportional to the square root of the temperature, doubling the rms speed would require increasing the temperature by a factor of 4. The ratio of the highest temperature ever recorded to the lowest temperature ever recorded is less than this:

$$\frac{T_2}{T_1} = \frac{331 \text{ K}}{144 \text{ K}} = 2.3$$

The temperature at the earth's surface is never high enough that nitrogen molecules move at twice the computed speed.

ASSESS We can use the square-root relationship to assess our computed result for the molecular speed. Figure 12.4 shows an rms speed of 510 m/s for nitrogen molecules at $20^\circ C$, or 293 K. Temperature T_1 is approximately half of this, so we'd expect to compute a speed that is lower by about $1/\sqrt{2}$, which is what we found.

Pressure

Everyone has some sense of the concept of *pressure*. If you get a hole in your bicycle tire, the higher-pressure air inside comes squirting out. It's hard to get the lid off a vacuum-sealed jar because of the low pressure inside. But just what is pressure?

Let's take an atomic-scale view of pressure, defining it in terms of the motion of particles of a gas. Suppose we have a sample of gas in a container with rigid walls. As particles in the gas move around, they sometimes collide with and bounce off the walls, creating a force on the walls, as illustrated in FIGURE 12.5.

The force on a surface is proportional to the area. Doubling the surface area will double the number of collisions and thus double the force. Rather than the force itself, a more useful quantity is the force-to-area ratio F/A . This ratio is a property of the gas itself, independent of the surface, and it's what we call the gas **pressure**:

$$p = \frac{F}{A} \quad (12.9)$$

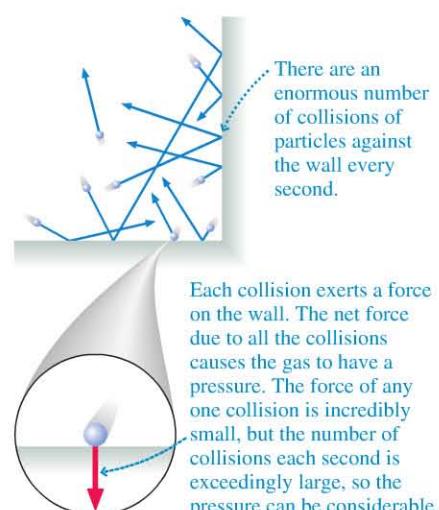
Definition of pressure in a gas

You can see from Equation 12.9 that a gas exerts a force of magnitude

$$F = pA \quad (12.10)$$

on a surface of area A . The force is *perpendicular* to the surface.

FIGURE 12.5 The pressure in a gas is due to the net force of the particles colliding with the walls.



NOTE ► Pressure itself is *not* a force, even though we sometimes talk informally about “the force exerted by the pressure.” The correct statement is that the *gas* exerts a force on a surface. ◀

From its definition, you can see that pressure has units of N/m^2 . The SI unit of pressure is the **pascal**, defined as

$$1 \text{ pascal} = 1 \text{ Pa} = 1 \frac{\text{N}}{\text{m}^2}$$

This unit is named for the 17th-century French scientist Blaise Pascal, who was one of the first to study gases. A pascal is a very small pressure, so we usually see pressures in kilopascals, where $1 \text{ kPa} = 1000 \text{ Pa}$.

The total force on the surface of your body due to the pressure of the atmosphere is over 40,000 pounds. Why doesn’t this enormous force pushing in simply crush you? The key is that there is also a force pushing out. **FIGURE 12.6a** shows an empty plastic soda bottle. The bottle isn’t a very sturdy structure, but the inward force due to the pressure of the atmosphere outside the bottle doesn’t crush it because there is an equal pressure due to air inside the bottle that pushes out. The forces pushing on both sides are quite large, but they exactly balance and so there is no net force.

A net pressure force is exerted only where there’s a pressure *difference* between the two sides of a surface. **FIGURE 12.6b** shows a surface of area A with a pressure difference $\Delta p = p_2 - p_1$ between the two sides. The net pressure force is

$$F_{\text{net}} = F_2 - F_1 = p_2 A - p_1 A = A(p_2 - p_1) = A \Delta p$$

This is the force that holds the lid on a vacuum-sealed jar, where the pressure inside is less than the pressure outside. To remove the lid, you have to exert a force greater than the force due to the pressure difference.

Decreasing the number of molecules in a container decreases the pressure because there are fewer collisions with the walls. The pressure in a completely empty container would be $p = 0 \text{ Pa}$. This is called a *perfect vacuum*. A perfect vacuum cannot be achieved because it’s impossible to remove every molecule from a region of space. In practice, a **vacuum** is an enclosed space in which $p \ll 1 \text{ atm}$. Using $p = 0 \text{ Pa}$ is then a good approximation.

Measuring Pressure

The most important pressure for our daily lives is the pressure of the atmosphere. The pressure of the atmosphere varies with altitude and the weather, but the global average pressure at sea level, called the *standard atmosphere*, is

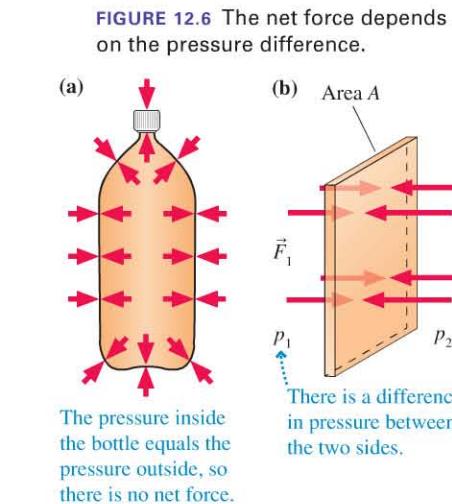
$$1 \text{ standard atmosphere} = 1 \text{ atm} = 101,300 \text{ Pa} = 101.3 \text{ kPa}$$

Just as we measured acceleration in units of g , we will often measure pressure in units of atm.

In the United States, pressure is often expressed in pounds per square inch, or psi. When you measure pressure in the tires on your car or bike, you probably use a gauge that reads in psi. The conversion factor is

$$1 \text{ atm} = 14.7 \text{ psi}$$

◀ **Too little pressure too fast** **BIO** This photo shows a rockfish, a popular game fish that can be caught at ocean depths up to 1000 ft. The great pressure at these depths is balanced by an equally large pressure inside the fish’s swim bladder, a gas-filled organ that we will learn more about in Chapter 13. If the fish is hooked and rapidly raised to the lower pressure of the surface, the pressure inside the swim bladder is suddenly much greater than the pressure outside, causing the swim bladder to expand dramatically, damaging nearby organs. When reeled in from a great depth, these fish seldom survive.



Because the effects of pressure depend on pressure differences, most gauges measure not the actual or *absolute pressure* p but what is called the **gauge pressure**. The gauge pressure, denoted p_g , is the pressure *in excess* of atmospheric pressure; that is,

$$p_g = p - 1 \text{ atm}$$

You need to add 1 atm to the reading of a pressure gauge to find the absolute pressure p you need for doing most calculations in this chapter.



A tire-pressure gauge reads the gauge pressure p_g , not the absolute pressure p . Zero gauge pressure doesn't mean that there's no pressure inside the tire. It means that the pressure inside the tire is equal to atmospheric pressure. There's no pressure difference between the inside and the outside, and so the tire is flat.

EXAMPLE 12.3 Finding the force due to a pressure difference

Patients suffering from decompression sickness may be treated in a hyperbaric oxygen chamber filled with oxygen at greater than atmospheric pressure. A cylindrical chamber with flat end plates of diameter 0.75 m is filled with oxygen to a gauge pressure of 27 kPa. What is the resulting force on the end plate of the cylinder?

PREPARE There is a force on the end plate because of the pressure *difference* between the inside and outside. 27 kPa is the pressure in excess of 1 atm. If we assume the pressure outside is 1 atm, then 27 kPa is Δp , the pressure difference across the surface.

SOLVE The end plate has area $A = \pi(0.75 \text{ m}/2)^2 = 0.442 \text{ m}^2$. The pressure difference results in a net force

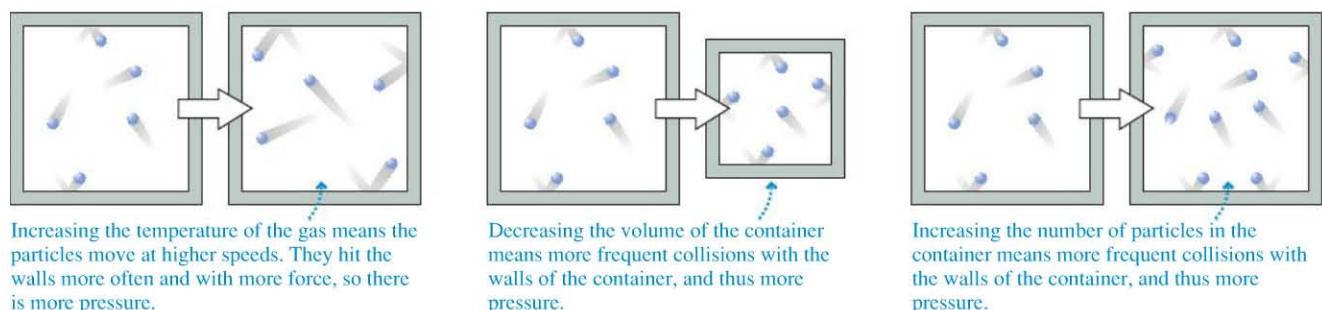
$$F_{\text{net}} = A \Delta p = (0.442 \text{ m}^2)(27,000 \text{ Pa}) = 12 \text{ kN}$$

ASSESS The area of the end plate is large, so we expect a large force. Our answer makes sense, although it is remarkable to think that this force results from the collisions of individual molecules with the plate. The large pressure force must be offset with an equally large force to keep the plate in place, so the end plate is fastened with stout bolts.

From Collisions to Pressure and the Ideal-Gas Law

We can use the fact that the pressure in a gas is due to the collisions of particles with the walls to make some qualitative predictions. FIGURE 12.7 presents a few such predictions.

FIGURE 12.7 Relating gas pressure to other variables.



Based on the reasoning in Figure 12.7, we expect the following proportionalities:

- Pressure should be proportional to the temperature of the gas: $p \propto T$.
- Pressure should be inversely proportional to the volume of the container: $p \propto 1/V$.
- Pressure should be proportional to the number of gas particles: $p \propto N$.

In fact, careful experiments back up each of these predictions. It's not hard to do a quantitative extension of the qualitative treatment above that looks at the collisions with the walls and the resulting forces. This theoretical analysis leads to the same



Pumping up a bicycle tire adds more particles to the fixed volume of the tire, increasing the pressure.

proportionalities noted above, and ultimately produces a single equation that expresses these proportionalities:

$$p = C \frac{NT}{V}$$

The proportionality constant C is none other than Boltzmann's constant k_B , which allows us to write

$$pV = Nk_B T \quad (12.11)$$

Ideal-gas law, version 1

Equation 12.11 is known as the **ideal-gas law**.

NOTE ► When using the ideal-gas law, make sure that pressure p is the *absolute* pressure (in Pa) and temperature T is the *absolute* temperature (in K). ◀

The important point of the above discussion is that the ideal-gas law is more than just an empirical finding; it can be understood on the basis of the atomic model and an analysis of the forces in the collisions of particles in a gas with the walls of their container.

Equation 12.11 is written in terms of the number N of particles in the gas, whereas the ideal-gas law is stated in chemistry in terms of the number n of moles. But the change is easy to make. The number of particles is $N = nN_A$, so we can rewrite Equation 12.11 as

$$pV = nN_A k_B T = nRT \quad (12.12)$$

Ideal-gas law, version 2

In this version of the equation, the proportionality constant—known as the *gas constant*—is

$$R = N_A k_B = 8.31 \text{ J/mol} \cdot \text{K}$$

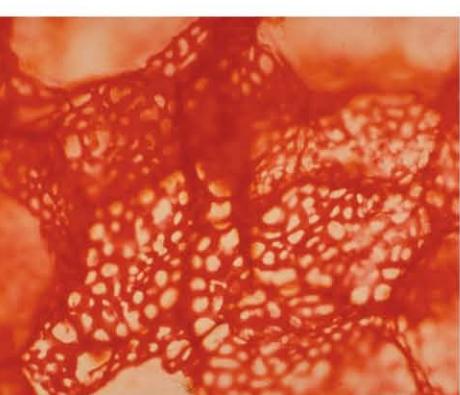
The units may seem unusual, but the product of Pa and m³, the units of pV , is equivalent to J.

NOTE ► You may have learned to work gas problems using units of atmospheres and liters. To do so, you had a different numerical value of R that was expressed in those units. In physics we always work gas problems in SI units, so the above value of R is the one you need to use. ◀

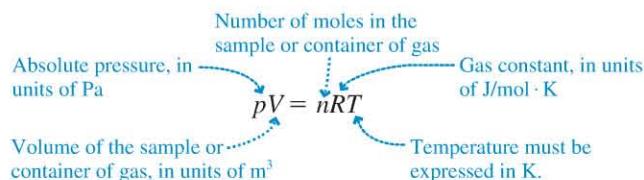
We will go back and forth between solving problems in terms of moles and numbers of atoms, so the following equality will be useful:

$$Nk_B = nN_A k_B = nR \quad (12.13)$$

◀ **Gas exchange in the lungs** **BIO** When you draw a breath, how does the oxygen get into your bloodstream? The atomic model provides some insight. The picture is a highly magnified view of the alveoli, air sacs in the lungs, surrounded by capillaries, fine blood vessels. The thin membranes of the alveoli and the capillaries are permeable—small molecules can move across them. If a permeable membrane separates two regions of space having different concentrations of a molecule, the rapid motion of the molecules causes a net transport of molecules in the direction of lower concentration, as we might expect from the discussion of why heat flows from hot to cold in Chapter 11. This transport—entirely due to the motion of the molecules—is known as **diffusion**. In the lungs, the higher concentration of oxygen in the alveoli drives the diffusion of oxygen into the blood in the capillaries. At the same time, carbon dioxide diffuses from the blood into the alveoli. Diffusion is a rapid and effective means of transport in this case because the membranes are thin. Diffusion won't work to get oxygen from the lungs to other parts of the body, because diffusion is slow over large distances, so the oxygenated blood must be pumped throughout the body.



Let's review the meanings and the units of the various quantities in the ideal-gas law:



EXAMPLE 12.4 Finding the volume of a mole

What volume is occupied by 1 mole of an ideal gas at a pressure of 1.00 atm and a temperature of 0°C?

PREPARE The first step in ideal-gas law calculations is to convert all quantities to SI units:

$$p = 1.00 \text{ atm} = 101.3 \times 10^3 \text{ Pa}$$

$$T = 0 + 273 = 273 \text{ K}$$

SOLVE We use the ideal-gas law equation to compute

$$V = \frac{nRT}{p} = \frac{(1.00 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})(273 \text{ K})}{101.3 \times 10^3 \text{ Pa}} = 0.0224 \text{ m}^3$$

We recall from earlier in the chapter that $1.00 \text{ m}^3 = 1000 \text{ L}$, so we can write

$$V = 22.4 \text{ L}$$

ASSESS At this temperature and pressure, we find that the volume of 1 mole of a gas is 22.4 L, a result you might recall from chemistry. When we do calculations using gases, it will be useful to keep this volume in mind to see if our answers make physical sense.

NOTE ► The conditions of this example, 1 atm pressure and 0°C, are called *standard temperature and pressure*, abbreviated **STP**. They are a common reference point for many gas processes. ◀

STOP TO THINK 12.2 A sample of ideal gas is in a sealed container. The temperature of the gas and the volume of the container are both increased. What other properties of the gas necessarily change? (More than one answer may be correct.)

- A. The rms speed of the gas atoms
- B. The thermal energy of the gas
- C. The pressure of the gas
- D. The number of molecules of gas

12.3 Ideal-Gas Processes

This chapter is about the thermal properties of matter. What changes occur to matter as you change the temperature? We can use the ideal-gas law to make such deductions for gases because it provides a connection among the pressure, volume, and temperature of a gas. For example, suppose you measure the pressure in the tires on your car on a cold morning. How much will the tire pressure increase after the tires warm up in the sun and the temperature of the air in the tires has increased? We will solve this problem later in the chapter, but, for now, note these properties of this process:

- The quantity of gas is fixed. No air is added to or removed from the tire. All of the processes we will consider will involve a fixed quantity of gas.
- There is a well-defined initial state. The initial values of pressure, volume, and temperature will be designated p_i , V_i , and T_i .
- There is a well-defined final state in which the pressure, volume, and temperature have values p_f , V_f , and T_f .

For gases in sealed containers, the number of moles (and the number of molecules) does not change. In that case, the ideal-gas law can be written as

$$\frac{pV}{T} = nR = \text{constant}$$

The values of the variables in the initial and final states are then related by

$$\frac{p_f V_f}{T_f} = \frac{p_i V_i}{T_i} \quad (12.14)$$

Initial and final states for an ideal gas in a sealed container

This before-and-after relationship between the two states, reminiscent of a conservation law, will be valuable for many problems.

NOTE ▶ Because pressure and volume appear on both sides of the equation, this is a rare case of an equation for which we may use any units we wish, not just SI units. However, temperature *must* be in K. Unit-conversion factors for pressure and volume are multiplicative factors, and the same factor on both sides of the equation cancels. But the conversion from K to °C is an *additive* factor, and additive factors in the denominator don't cancel. ◀

pV Diagrams

It's useful to represent ideal-gas processes on a graph called a **pV diagram**. The important idea behind a pV diagram is that each point on the graph represents a single, unique state of the gas. This may seem surprising because a point on a graph specifies only the value of pressure and volume. But knowing p and V , and assuming that n is known for a sealed container, we can find the temperature from the ideal-gas law. Thus each point on a pV diagram actually represents a triplet of values (p , V , T) specifying the state of the gas.

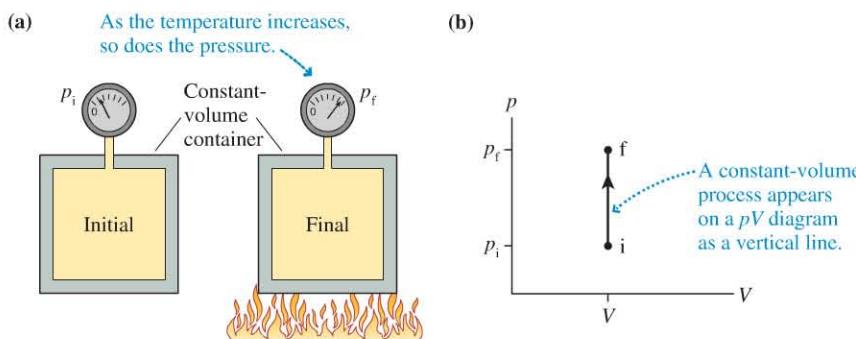
For example, FIGURE 12.8a is a pV diagram showing three states of a system consisting of 1 mol of gas. The values of p and V can be read from the axes, then the temperature at that point calculated from the ideal-gas law. An ideal-gas process—a process that changes the state of the gas by, for example, heating it or compressing it—can be represented as a “trajectory” in the pV diagram. FIGURE 12.8b shows one possible process by which the gas of Figure 12.8a is changed from state 1 to state 3.

Constant-Volume Processes

Suppose you have a gas in the closed, rigid container shown in FIGURE 12.9a. Warming the gas will raise its pressure without changing its volume. This is an example of a **constant-volume process**. $V_f = V_i$ for a constant-volume process.

Because the value of V doesn't change, this process is shown as the vertical line $i \rightarrow f$ on the pV diagram of FIGURE 12.9b. A **constant-volume process appears on a pV diagram as a vertical line**.

FIGURE 12.9 A constant-volume process.



EXAMPLE 12.5 Computing tire pressure on a hot day

The pressure in a car tire is 30.0 psi on a cool morning when the air temperature is 0°C. After the day warms up and bright sun shines on the black tire, the temperature of the air inside the tire reaches 30°C. What is the tire pressure at this temperature?

PREPARE A tire is (to a good approximation) a sealed container with constant volume; this is a constant-volume process. The measured tire pressure is a gauge pressure, but the ideal-gas law requires an absolute pressure. We must correct for this. The initial pressure is

$$p_i = (p_g)_i + 1.00 \text{ atm} = 30.0 \text{ psi} + 14.7 \text{ psi} = 44.7 \text{ psi}$$

Temperatures must be in kelvin, so we convert:

$$T_i = 0^\circ\text{C} + 273 = 273 \text{ K}$$

$$T_f = 30^\circ\text{C} + 273 = 303 \text{ K}$$

SOLVE The gas is in a sealed container, so we can use the ideal-gas law as given in Equation 12.14 to solve for the final pressure. In this equation, we divide both sides by V_f , and then cancel the ratio of the two volumes, which is equal to 1 for this constant-volume process:

$$p_f = p_i \frac{V_i}{V_f} \frac{T_f}{T_i} = p_i \frac{T_f}{T_i}$$

The units for p_f will be the same as those for p_i , so we can keep the initial pressure in psi. The pressure at the higher temperature is

$$p_f = 44.7 \text{ psi} \times \frac{303 \text{ K}}{273 \text{ K}} = 49.6 \text{ psi}$$

This is an absolute pressure, but the problem asks for the measured pressure in the tire—a gauge pressure. Converting to gauge pressure gives

$$(p_g)_f = p_f - 1.00 \text{ atm} = 49.6 \text{ psi} - 14.7 \text{ psi} = 34.9 \text{ psi}$$

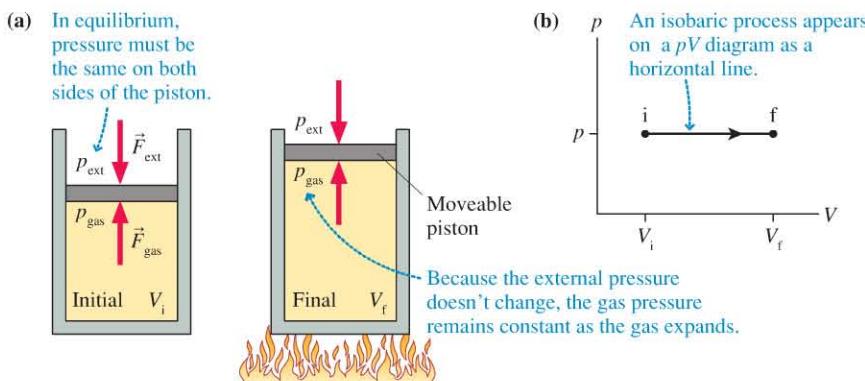
ASSESS The temperature has changed by 30 K, which is a bit more than 10% of the initial temperature, so we expect a large change in pressure. Our result seems reasonable, and it has practical implications: If you check the pressure in your tires when they are at a particular temperature, don't expect the pressure to be the same when conditions change!

Constant-Pressure Processes

Many gas processes take place at a constant, unchanging pressure. A constant-pressure process is also called an **isobaric process**. $p_f = p_i$ for an isobaric process.

One way to produce an isobaric process is shown in **FIGURE 12.10a**, where a gas is sealed in a cylinder by a lightweight, tight-fitting cap—a *piston*—that is free to slide up and down. In fact, the piston *will* slide up or down, compressing or expanding the gas inside, until it reaches the position at which $p_{\text{gas}} = p_{\text{ext}}$. That's the equilibrium position for the piston, the position at which the upward force $F_{\text{gas}} = p_{\text{gas}}A$, where A is the area of the face of the piston, exactly balances the downward force $F_{\text{ext}} = p_{\text{ext}}A$ due to the external pressure p_{ext} . The gas pressure in this situation is controlled by the external pressure. As long as the external pressure doesn't change, neither can the gas pressure inside the cylinder.

FIGURE 12.10 A constant-pressure (isobaric) process.



Suppose we heat the gas in the cylinder. The gas pressure doesn't change because the pressure is controlled by the unchanging external pressure, not by the temperature. But as the temperature rises, the faster-moving atoms cause the gas to expand, pushing the piston outward. Because the pressure is always the same, this is an isobaric process, with a trajectory as shown in **FIGURE 12.10b**. A **constant-pressure process** appears on a pV diagram as a horizontal line.

EXAMPLE 12.6 A constant-pressure compression

A gas in a cylinder with a movable piston occupies 50.0 cm^3 at 50°C . The gas is cooled at constant pressure until the temperature is 10°C . What is the final volume?

PREPARE This is a sealed container, so we can use Equation 12.14. The pressure of the gas doesn't change, so this is an isobaric process with $p_i/p_f = 1$.

The temperatures must be in kelvin, so we convert:

$$T_i = 50^\circ\text{C} + 273 = 323 \text{ K}$$

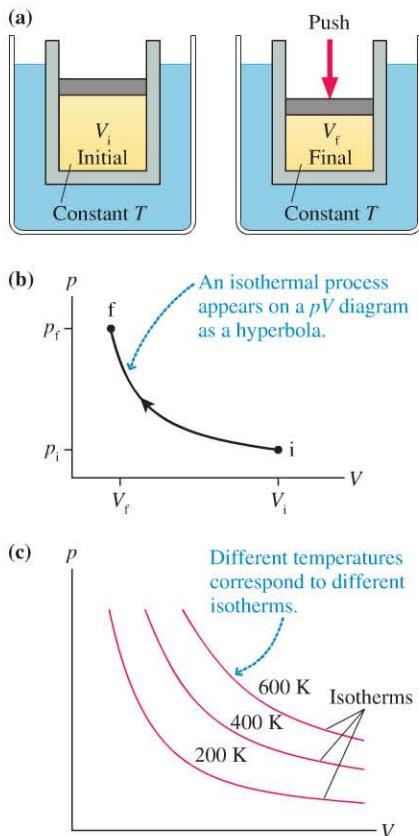
$$T_f = 10^\circ\text{C} + 273 = 283 \text{ K}$$

SOLVE We can use the ideal-gas law for a sealed container to solve for V_f :

$$V_f = V_i \frac{p_i}{p_f} \frac{T_f}{T_i} = 50.0 \text{ cm}^3 \times 1 \times \frac{283 \text{ K}}{323 \text{ K}} = 43.8 \text{ cm}^3$$

ASSESS In this example and the previous one, we have not converted pressure and volume units because these multiplicative factors cancel. But we did convert temperature to kelvin because this *additive* factor does *not* cancel.

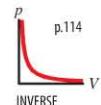
FIGURE 12.11 A constant-temperature (isothermal) process.

**Constant-Temperature Processes**

A constant-temperature process is also called an **isothermal process**. For an isothermal process $T_f = T_i$. One possible isothermal process is illustrated in **FIGURE 12.11**. A piston is being pushed down to compress a gas, but the gas cylinder is submerged in a large container of liquid that is held at a constant temperature. If the piston is pushed *slowly*, then heat-energy transfer through the walls of the cylinder will keep the gas at the same temperature as the surrounding liquid. This would be an *isothermal compression*. The reverse process, with the piston slowly pulled out, would be an *isothermal expansion*.

Representing an isothermal process on the pV diagram is a little more complicated than the two preceding processes because both p and V change. As long as T remains fixed, we have the relationship

$$p = \frac{nRT}{V} = \frac{\text{constant}}{V} \quad (12.15)$$



Because there is an inverse relationship between p and V , the graph of an isothermal process is a *hyperbola*.

The process shown as $i \rightarrow f$ in Figure 12.11b represents the *isothermal compression* shown in Figure 12.11a. An *isothermal expansion* would move in the opposite direction along the hyperbola. The graph of an isothermal process is known as an **isotherm**.

The location of the hyperbola depends on the value of T . If we use a higher constant temperature for the process in Figure 12.11a, the isotherm will move farther from the origin of the pV diagram. Figure 12.11c shows three isotherms for this process at three different temperatures. A gas undergoing an isothermal process will move along the isotherm for the appropriate temperature.

EXAMPLE 12.7 Compressing air in the lungs

A snorkeler takes a deep breath at the surface, filling his lungs with 4.0 L of air. He then descends to a depth of 5.0 m , where the pressure is 0.50 atm higher than at the surface. At this depth, what is the volume of air in the snorkeler's lungs?

PREPARE At the surface, the pressure of the air inside the snorkeler's lungs is 1.0 atm —it's atmospheric pressure at sea level. As he descends, the pressure inside his lungs must rise to match the pressure of the surrounding water, because the body can't sustain large pressure differences between inside and out. Further, the air stays at body temperature, making this an isothermal process with $T_f = T_i$.

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The temperatures must be in kelvin, so we convert:

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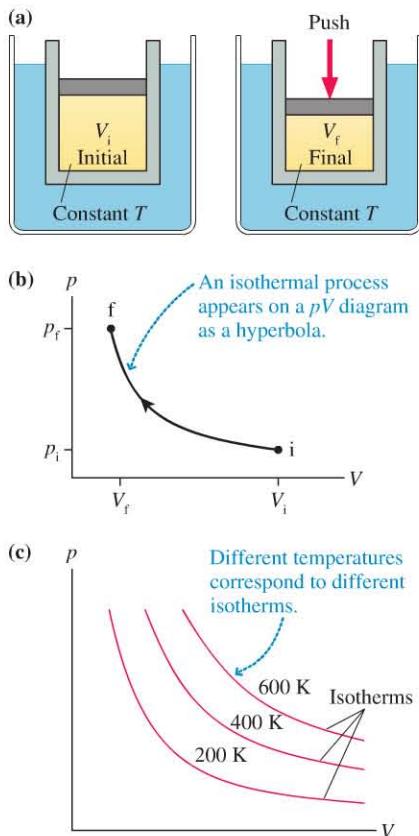
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SOLVE We can use the ideal-gas law for a sealed container to solve for V_f :

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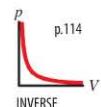
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Although we've shown this result only for an isobaric process, it turns out to be true for all ideal-gas processes. That is, as **FIGURE 12.13b** shows,

$$W_{\text{gas}} = \text{area under the } pV \text{ graph between } V_i \text{ and } V_f$$

There are a few things to clarify:

- In order for the gas to do work, its volume must change. No work is done in a constant-volume process.
- The simple relationship of Equation 12.16 applies only to constant-pressure processes. For any other ideal-gas process, you must use the geometry of the pV diagram to calculate the area under the graph.
- To calculate work, pressure must be in Pa and volume in m^3 . The product of Pa (which is N/m^2) and m^3 is $\text{N} \cdot \text{m}$. But $1 \text{ N} \cdot \text{m}$ is 1 J—the unit of work and energy.
- W_{gas} is positive if the gas expands ($\Delta V > 0$). The gas does work by pushing against the piston. In this case, the work done is energy transferred out of the system, and the energy of the gas decreases. W_{gas} is negative if the piston compresses the gas ($\Delta V < 0$) because the force \vec{F}_{gas} is opposite the displacement of the piston. Energy is transferred into the system as work, and the energy of the gas increases. We often say "work is done *on* the gas," but this just means that W_{gas} is negative.

In the first law of thermodynamics, $\Delta E_{\text{th}} = Q + W$, W is the work done by the environment—that is, by force \vec{F}_{ext} acting on the system. But \vec{F}_{ext} and \vec{F}_{gas} are equal and opposite forces, as we noted above, so the work done by the environment is the negative of the work done by the gas: $W = -W_{\text{gas}}$. Consequently, the first law of thermodynamics can be written as

$$\Delta E_{\text{th}} = Q - W_{\text{gas}} \quad (12.17)$$

In Chapter 11 we learned that the thermal energy of an ideal gas depends only on its temperature as $E_{\text{th}} = \frac{3}{2}Nk_B T$. Equation 12.13 tells us that Nk_B is equal to nR , so we can write the change in thermal energy of an ideal gas as

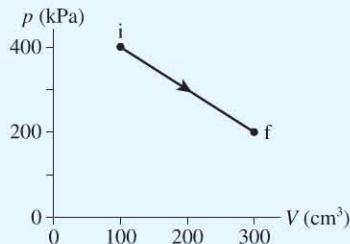
$$\Delta E_{\text{th}} = \frac{3}{2}Nk_B \Delta T = \frac{3}{2}nR \Delta T \quad (12.18)$$

EXAMPLE 12.8

Thermodynamics of an expanding gas

A cylinder with a movable piston contains 0.016 mol of helium. A researcher expands the gas via the process illustrated in **FIGURE 12.14**. To achieve this, does she need to heat the gas? If so, how much heat energy must be added or removed?

FIGURE 12.14 pV diagram for Example 12.8.



PREPARE As the gas expands, it does work on the piston. Its temperature may change as well, implying a change in thermal energy. We can use the version of the first law of thermodynamics in Equation 12.17 to describe the energy changes in the gas. We'll first find the change in thermal energy by computing the temperature change, then calculate how much work is done by looking at

the area under the graph. Once we know W_{gas} and ΔE_{th} , we can use the first law to determine the sign and the magnitude of the heat—telling us whether heat energy goes in or out, and how much.

The graph tells us the pressure and the volume, so we can use the ideal-gas law to compute the temperature at the initial and final points. To do this we'll need the volumes in SI units. Reading the initial and final volumes on the graph and converting, we find

$$V_i = 100 \text{ cm}^3 \times \frac{1 \text{ m}^3}{10^6 \text{ cm}^3} = 1.0 \times 10^{-4} \text{ m}^3$$

$$V_f = 300 \text{ cm}^3 \times \frac{1 \text{ m}^3}{10^6 \text{ cm}^3} = 3.0 \times 10^{-4} \text{ m}^3$$

SOLVE The initial and final temperatures are found using the ideal-gas law:

$$T_i = \frac{p_i V_i}{nR} = \frac{(4.0 \times 10^5 \text{ Pa})(1.0 \times 10^{-4} \text{ m}^3)}{(0.016 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})} = 300 \text{ K}$$

$$T_f = \frac{p_f V_f}{nR} = \frac{(2.0 \times 10^5 \text{ Pa})(3.0 \times 10^{-4} \text{ m}^3)}{(0.016 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})} = 450 \text{ K}$$

The temperature increases, and so must the thermal energy. We can use Equation 12.18 to compute this change:

$$\Delta E_{\text{th}} = \frac{3}{2}(0.016 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})(450 \text{ K} - 300 \text{ K}) = 30 \text{ J}$$

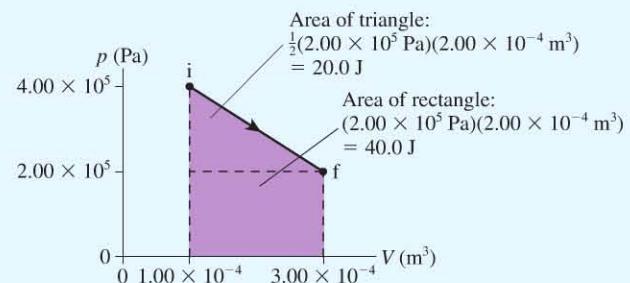
The other piece of the puzzle is to compute the work done. We do this by finding the area under the graph for the process. **FIGURE 12.15** shows that we can do this calculation by viewing the area as a triangle on top of a rectangle. Notice that the areas are in joules because they are the product of Pa and m³. The total work is

$$\begin{aligned} W_{\text{gas}} &= \text{area of triangle} + \text{area of rectangle} \\ &= 20 \text{ J} + 40 \text{ J} = 60 \text{ J} \end{aligned}$$

Now we can use the first law as written in Equation 12.17 to find the heat:

$$Q = \Delta E_{\text{th}} + W_{\text{gas}} = 30 \text{ J} + 60 \text{ J} = 90 \text{ J}$$

FIGURE 12.15 The work done by the expanding gas is the total area under the graph.



This is a positive number, so—using the conventions introduced in Chapter 11—we see that 90 J of heat energy must be added to the gas.

ASSESS The gas does work—a loss of energy—but its temperature increases, so it makes sense that heat energy must be added.

Adiabatic Processes

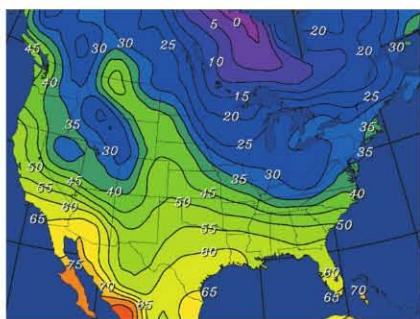
You may have noticed that when you pump up a bicycle tire with a hand pump, the pump gets warm. We noted the reason for this in Chapter 11: When you press down on the handle of the pump, a piston in the pump's chamber compresses the gas. According to the first law of thermodynamics, doing work on the gas increases its thermal energy. So the gas temperature goes up, and heat is then transferred through the walls of the pump to your hand.

Now suppose you compress a gas in an insulated container, so that no heat is exchanged with the environment, or you compress a gas so quickly that there is no time for heat to be transferred. In either case, $Q = 0$. If a gas process has $Q = 0$, for either a compression or an expansion, we call this an **adiabatic process**.

An expanding gas does work, so $W_{\text{gas}} > 0$. If the expansion is adiabatic, meaning $Q = 0$, then the first law of thermodynamics as written in Equation 12.17 tells us that $\Delta E_{\text{th}} < 0$. Temperature is proportional to thermal energy, so the temperature will decrease as well. **An adiabatic expansion lowers the temperature of a gas.** If the gas is compressed, work is done on the gas ($W_{\text{gas}} < 0$). If the compression is adiabatic, the first law of thermodynamics implies that $\Delta E_{\text{th}} > 0$ and thus that the temperature increases. **An adiabatic compression raises the temperature of a gas.** Adiabatic processes allow you to use work, rather than heat, to change the temperature of a gas.

Adiabatic processes have many important applications. In the atmosphere, large masses of air moving across the planet exchange heat energy with their surroundings very slowly, so adiabatic expansions and compressions can lead to large and abrupt changes in temperature, as we see in the discussion of the Chinook wind.

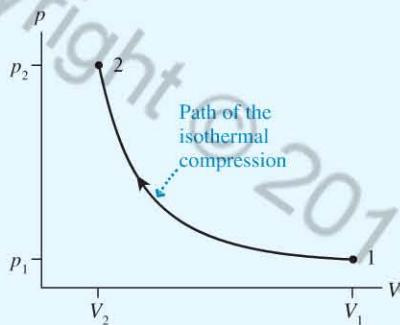
► **Warm mountain winds** This image shows surface temperatures (in °F) in North America on a winter day. Notice the bright green area of unseasonably warm temperatures extending north and west from the center of the continent. On this day, a strong westerly wind, known as a Chinook wind, was blowing down off the Rocky Mountains, rapidly moving from high elevations (and low pressures) to low elevations (and higher pressures). The air was rapidly compressed as it descended. The compression was so rapid that no heat was exchanged with the environment, so this was an adiabatic process that significantly increased the air temperature.



CONCEPTUAL EXAMPLE 12.9**Adiabatic curves on a pV diagram**

FIGURE 12.16 shows the pV diagram of a gas undergoing an isothermal compression from point 1 to point 2. Sketch how the pV diagram would look if the gas were compressed from point 1 to the same final pressure by a rapid adiabatic compression.

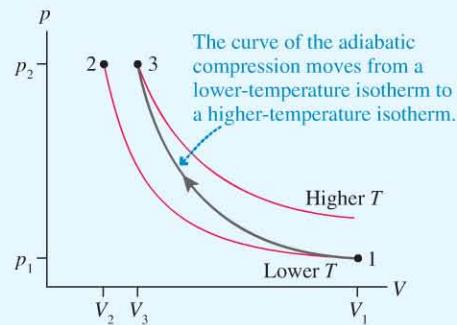
FIGURE 12.16 pV diagram for an isothermal compression.



REASON An adiabatic compression increases the temperature of the gas as the work done on the gas is transformed into thermal energy. Consequently, as seen in **FIGURE 12.17**, the curve of the

adiabatic compression cuts across the isotherms to end on a higher-temperature isotherm when the gas pressure reaches p_2 .

FIGURE 12.17 pV diagram for an adiabatic compression.



ASSESS In an isothermal compression, heat energy is transferred out of the gas so that the gas temperature stays the same. This heat transfer doesn't happen in an adiabatic compression, so we'd expect the gas to have a higher final temperature. In general, the temperature at the final point of an adiabatic compression is higher than at the starting point. Similarly, an adiabatic expansion ends on a lower-temperature isotherm.

STOP TO THINK 12.3

What is the ratio T_f/T_i for this process?

- A. 1/4
- B. 1/2
- C. 1 (no change)
- D. 2
- E. 4
- F. There is not enough information to decide.

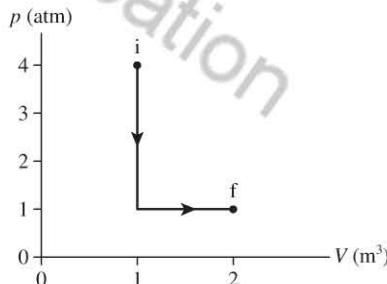
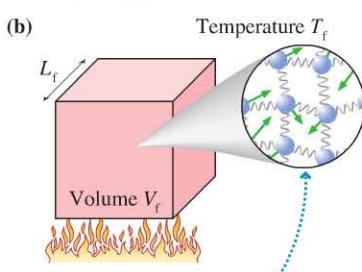
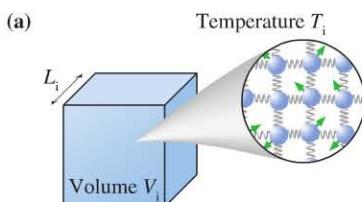


FIGURE 12.18 Increasing the temperature of a solid causes it to expand.



If an object is heated, its thermal energy increases. The atoms and molecules jiggle around at higher speeds and move farther apart.

12.4 Thermal Expansion

The bonds between atoms in solids and liquids mean that solids and liquids are much less compressible than gases, as we've noted. But raising the temperature of a solid or a liquid does produce a small, measurable change in volume. You've seen this principle at work if you've ever used a thermometer that contains mercury or alcohol (the kind with the red liquid) to measure a temperature. When you immerse the bulb of the thermometer in something warm, the liquid moves up the column for a simple reason: It expands. This **thermal expansion** underlies many practical phenomena.

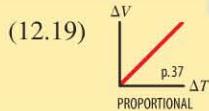
FIGURE 12.18a shows a cube of material, initially at temperature T_i . The edge of the cube has length L_i , and the cube's volume is V_i . We then heat the cube, increasing its temperature to T_f , as shown in **FIGURE 12.18b**. The atoms now jiggle around faster, so

the bonds are stretched. This increases the average distance between atoms, and so the cube's volume increases to V_f .

For most substances, the change in volume $\Delta V = V_f - V_i$ is linearly related to the change in temperature $\Delta T = T_f - T_i$ by

$$\Delta V = \beta V_i \Delta T$$

Volume thermal expansion



The constant β is known as the **coefficient of volume expansion**. Its value depends on the nature of the bonds in a material; it can vary quite a bit from one substance to another. Because ΔT is measured in K, the units of β are K^{-1} .

As a solid's volume increases, each of its linear dimensions increases as well. We can write a similar expression for this linear thermal expansion. If an object of initial length L_i undergoes a temperature change ΔT , its length changes to L_f . The change in length, $\Delta L = L_f - L_i$, is given by

$$\Delta L = \alpha L_i \Delta T \quad (12.20)$$

Linear thermal expansion

The constant α is the **coefficient of linear expansion**. Note that Equations 12.19 and 12.20 apply equally well to thermal *contractions*, in which case both ΔT and ΔV (or ΔL) are negative.

NOTE ► The expressions for thermal expansion are only approximate expressions that apply over a limited range of temperatures. They are what we call **empirical formulas**; they are a good fit to measured data, but they do not represent any underlying fundamental law. There are materials that do *not* closely follow Equations 12.19 and 12.20; water at low temperatures is one example, as we will see. ◀

Values of α and β for some common materials are listed in Table 12.3. The volume expansion of a liquid can be measured, but, because a liquid can change shape, we don't assign a coefficient of linear expansion to a liquid. Although α and β do vary slightly with temperature, the room-temperature values in Table 12.3 are adequate for all problems in this book. The values of α and β are very small, so the changes in length and volume, ΔL and ΔV , are a small fraction of the original length.



Expanding spans A long steel bridge will slightly increase in length on a hot day and decrease on a cold day. Thermal expansion joints let the bridge's length change without causing the roadway to buckle.

TABLE 12.3 Coefficients of linear and volume thermal expansion at 20°C

Substance	Linear α (K $^{-1}$)	Volume β (K $^{-1}$)
Aluminum	23×10^{-6}	69×10^{-6}
Glass	9×10^{-6}	27×10^{-6}
Iron or steel	12×10^{-6}	36×10^{-6}
Concrete	12×10^{-6}	36×10^{-6}
Ethyl alcohol		1100×10^{-6}
Water		210×10^{-6}
Air (and other gases)		3400×10^{-6}

EXAMPLE 12.10 How much closer to space?

The height of the Space Needle, a steel observation tower in Seattle, is 180 meters on a 0°C winter day. How much taller is it on a hot summer day when the temperature is 30°C?

PREPARE The steel expands because of an increase in temperature, which is

$$\Delta T = T_f - T_i = 30^\circ\text{C} - 0^\circ\text{C} = 30^\circ\text{C} = 30 \text{ K}$$

SOLVE The coefficient of linear expansion is given in Table 12.3; we can use this value in Equation 12.20 to compute the increase in height:

$$\Delta L = \alpha L_i \Delta T = (12 \times 10^{-6} \text{ K}^{-1})(180 \text{ m})(30 \text{ K}) = 0.065 \text{ m}$$

ASSESS You don't notice buildings getting taller on hot days, so we expect the final answer to be small. The change is a small fraction of the height of the tower, as we expect. Our answer makes physical sense. Compared to 180 m, an expansion of 6.5 cm is not something you would easily notice—but it isn't negligible. The thermal expansion of structural elements in towers and bridges must be accounted for in the design to avoid damaging stresses. When designers failed to properly account for thermal stresses in the marble panels cladding the Amoco Building in Chicago, all 43,000 panels had to be replaced, at great cost.

TRY IT YOURSELF

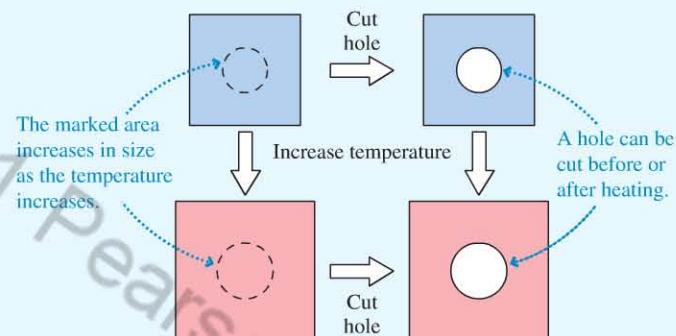
Thermal expansion to the rescue If you have a stubborn lid on a glass jar, try this: Put the lid under very hot water for a short time. Heating the lid and the jar makes them both expand, but, as you can see from the data in Table 12.3, the steel lid—and the opening in the lid that fits over the glass jar—expands by more than the glass jar. The jar lid is now looser and can be more easily removed.

CONCEPTUAL EXAMPLE 12.11 What happens to the hole?

A metal plate has a circular hole in it. As the plate is heated, does the hole get larger or smaller?

REASON As the plate expands, you might think the hole would shrink (with the metal expanding into the hole). But suppose we took a metal plate and simply drew a circle where a hole could be cut. On heating, the plate and the marked area both expand, as we see in **FIGURE 12.19**. We could cut a hole on the marked line before or after heating; the size of the hole would be larger in the latter case. Therefore, the size of the hole must expand as the plate expands.

FIGURE 12.19 The thermal expansion of a hole.

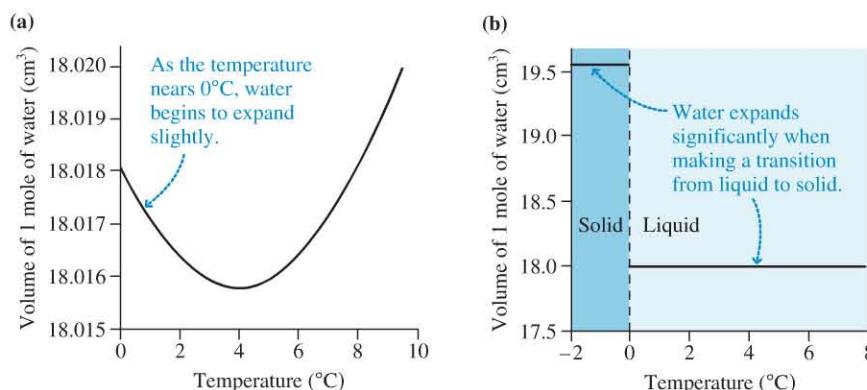


ASSESS This somewhat counterintuitive result has many practical applications, as you can see from the “Try It Yourself” exercise on loosening a jar lid.

Special Properties of Water and Ice

Water, the most important molecule for life, differs from other liquids in many important ways. If you cool a sample of water toward its freezing point of 0°C , you would expect its volume to decrease. **FIGURE 12.20a** shows a graph of the volume of a mole of water versus temperature. As the water is cooled, you can see that the volume indeed decreases—to a point. Once the water is cooled below 4°C , further cooling results in an *increase* in volume.

FIGURE 12.20 Variation of the volume of a mole of water with temperature.



This anomalous behavior of water is a result of the charge distribution on the water molecules, which determines how nearby molecules interact with each other. As water approaches the freezing point, molecules begin to form clusters that are more strongly bound. Water molecules must actually get a bit farther apart to form

such clusters, meaning the volume increases. Freezing the water results in an even greater increase in volume, as illustrated in **FIGURE 12.20b**.

This expansion on freezing has important consequences. In most materials, the solid phase is denser than the liquid phase, and so the solid material sinks. Because water *expands* as it freezes, ice is less dense than liquid water and floats. Not only do ice cubes float in your cold drink, but a lake freezes by forming ice on the top rather than at the bottom. This layer of ice then insulates the water below from much colder air above. Thus most lakes do not freeze solid, allowing aquatic life to survive even the harshest winters.

STOP TO THINK 12.4 An aluminum ring is tight around a solid iron rod. If we wish to loosen the ring to remove it from the rod, we should

- A. Increase the temperature of the ring and rod.
- B. Decrease the temperature of the ring and rod.

12.5 Specific Heat and Heat of Transformation

If you hold a glass of cold water in your hand, the heat from your hand will raise the temperature of the water. Similarly, the heat from your hand will melt an ice cube; this is what we will call a *phase change*. Increasing the temperature and changing the phase are two possible consequences of heating a system, and they form an important part of our understanding of the thermal properties of matter.

Specific Heat

Adding 4190 J of heat energy to 1 kg of water raises its temperature by 1 K. If you are fortunate enough to have 1 kg of gold, you need only 129 J of heat to raise its temperature by 1 K. The amount of heat that raises the temperature of 1 kg of a substance by 1 K is called the **specific heat** of that substance. The symbol for specific heat is c . Water has specific heat $c_{\text{water}} = 4190 \text{ J/kg} \cdot \text{K}$, and the specific heat of gold is $c_{\text{gold}} = 129 \text{ J/kg} \cdot \text{K}$. Specific heat depends only on the material from which an object is made. Table 12.4 lists the specific heats for some common liquids and solids.

If heat c is required to raise the temperature of 1 kg of a substance by 1 K, then heat Mc is needed to raise the temperature of mass M by 1 K and $Mc \Delta T$ is needed to raise the temperature of mass M by ΔT . In general, the heat needed to bring about a temperature change ΔT is

$$Q = Mc \Delta T \quad (12.21)$$

Heat needed to produce a temperature change ΔT for mass M with specific heat c

TABLE 12.4 Specific heats of solids and liquids

Substance	$c (\text{J/kg} \cdot \text{K})$
Solids	
Lead	128
Gold	129
Copper	385
Iron	449
Aluminum	900
Water ice	2090
Mammalian body	3400
Liquids	
Mercury	140
Ethyl alcohol	2400
Water	4190

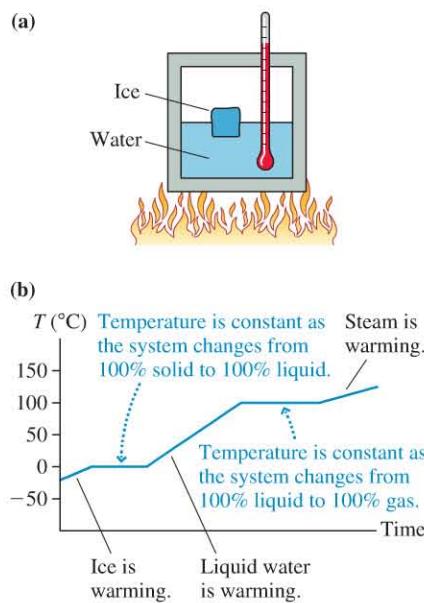
Q can be either positive (temperature goes up) or negative (temperature goes down).

Because $\Delta T = Q/Mc$, it takes more heat energy to change the temperature of a substance with a large specific heat than to change the temperature of a substance with a small specific heat. Water, with a very large specific heat, is slow to warm up and slow to cool down. This is fortunate for us. The large “thermal inertia” of water is essential for the biological processes of life.



Hurricane season At night, the large specific heat of water prevents the temperature of a body of water from dropping nearly as much as that of the surrounding air. Early in the morning, water vapor evaporating from a warm lake quickly condenses in the colder air above, forming mist. During the day, the opposite happens: The air becomes much warmer than the water. This lag between air and water temperatures is also at work at larger scales. Ocean temperatures in the northern hemisphere don't reach their maximum values until the late summer or early fall. This is why hurricanes, which are fueled by the energy in warm ocean water, happen in the fall but not the spring.

FIGURE 12.21 The temperature as a function of time as water is transformed from solid to liquid to gas.



EXAMPLE 12.12 Energy to run a fever

A 70 kg student catches the flu, and his body temperature increases from 37.0°C (98.6°F) to 39.0°C (102.2°F). How much energy is required to raise his body's temperature?

PREPARE The increase in temperature requires the addition of energy. The change in temperature ΔT is 2.0°C , or 2.0 K .

SOLVE Raising the temperature of the body uses energy supplied internally from the chemical reactions of the body's metabolism, which transfer heat to the body. The specific heat of the body is given in Table 12.4 as $3400\text{ J/kg} \cdot \text{K}$. We can use Equation 12.21 to find the necessary heat energy:

$$Q = Mc \Delta T = (70\text{ kg})(3400\text{ J/kg} \cdot \text{K})(2.0\text{ K}) = 4.8 \times 10^5\text{ J}$$

ASSESS The body is mostly water, with a large specific heat, and the mass of the body is large, so we'd expect a large amount of energy to be necessary. Looking back to Chapter 11, we see that this is approximately the energy in a large apple, or the amount of energy required to walk 1 mile.

Phase Changes

The temperature inside the freezer compartment of a refrigerator is typically about -20°C . Suppose you remove a few ice cubes from the freezer and place them in a sealed container with a thermometer. Then, as in **FIGURE 12.21a**, you put a steady flame under the container. We'll assume that the heating is done slowly so that the inside of the container always has a single, uniform temperature.

FIGURE 12.21b shows the temperature as a function of time. After steadily rising from the initial -20°C , the temperature remains fixed at 0°C for an extended period of time. This is the interval of time during which the ice melts. As it's melting, the ice temperature is 0°C and the liquid water temperature is 0°C . Even though the system is being heated, the temperature doesn't begin to rise until all the ice has melted.

The atomic model can give us some insight into what is taking place. The thermal energy of a solid is the kinetic energy of the vibrating atoms plus the potential energy of stretched and compressed molecular bonds. If you heat a solid, at some point the thermal energy gets so large that the molecular bonds begin to break, allowing the atoms to move around—the solid begins to melt. Continued heating will produce further melting, as the energy is used to break more bonds. The temperature will not rise until all of the bonds are broken. If you wish to return the liquid to a solid, this energy must be removed.

NOTE ► In everyday language, the three phases of water are called *ice*, *water*, and *steam*. The term “water” implies the liquid phase. Scientifically, these are the solid, liquid, and gas phases of the compound called *water*. When we are working with different phases of water, we'll use the term “water” in the scientific sense of a collection of H_2O molecules. We'll say either *liquid* or *liquid water* to denote the liquid phase. ◀

CONCEPTUAL EXAMPLE 12.13 Strategy for cooling a drink

If you have a warm soda that you wish to cool, is it more effective to add 25 g of liquid water at 0°C or 25 g of water ice at 0°C ?

REASON If you add liquid water at 0°C , heat will be transferred from the soda to the water, raising the temperature of the water and lowering that of the soda. If you add water ice at 0°C , heat first will be transferred from the soda to the ice to melt it, transforming the 0°C ice to 0°C liquid water, then will be transferred to the liquid water to raise its temperature. Thus more thermal energy will be removed from the soda, giving it a lower final temperature, if ice is used rather than liquid water.

ASSESS This makes sense because you know that this is what you do in practice. To cool a drink, you drop in an ice cube.

The temperature at which a solid becomes a liquid or, if the thermal energy is reduced, a liquid becomes a solid is called the **melting point** or the **freezing point**. Melting and freezing are *phase changes*. A system at the melting point is in **phase equilibrium**, meaning that any amount of solid can coexist with any amount of liquid. Raise the temperature ever so slightly and the entire system soon becomes liquid. Lower it slightly and it all becomes solid.

You can see another region of phase equilibrium of water in Figure 12.21b at 100°C. This is a phase equilibrium between the liquid phase and the gas phase, and any amount of liquid can coexist with any amount of gas at this temperature. As heat is added to the system, the temperature stays the same. The added energy is used to break bonds between the liquid molecules, allowing them to move into the gas phase. The temperature at which a gas becomes a liquid or a liquid becomes a gas is called the **condensation point** or the **boiling point**.

NOTE ► Liquid water becomes solid ice at 0°C, but that doesn't mean the temperature of ice is always 0°C. Ice reaches the temperature of its surroundings. If the air temperature in a freezer is –20°C, then the ice temperature is –20°C. Likewise, steam can be heated to temperatures above 100°C. That doesn't happen when you boil water on the stove because the steam escapes, but steam can be heated far above 100°C in a sealed container. ◀

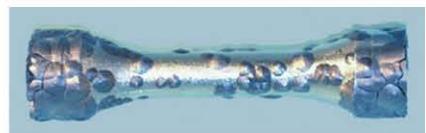
CONCEPTUAL EXAMPLE 12.14 Fast or slow boil?

You are cooking pasta on the stove; the water is at a slow boil. Will the pasta cook more quickly if you turn up the burner on the stove so that the water is at a fast boil?

REASON Water boils at 100°C; no matter how vigorously the water is boiling, the temperature is the same. It is the temperature of the water that determines how fast the cooking takes place.

Adding heat at a faster rate will make the water boil away more rapidly but will not change the temperature—and will not alter the cooking time.

ASSESS This result may seem counterintuitive, but you can try the experiment next time you cook pasta!



Tin pest Some substances have multiple solid phases in which the atoms are arranged in lattices with different structures. Such substances can undergo a phase change between the solid phases; the atoms rearrange themselves, and the properties of the substance can change. The element tin slowly changes from the usual “white” phase to a brittle “gray” phase at low temperatures. The tin sample in this photo was chilled for one year; the resulting “tin pest” is quite obvious.

Some historians speculate that this phase change destroyed the tin buttons on the coats of Napoleon's troops during their winter invasion of Russia in 1812, contributing to their defeat as they froze in the bitter cold.

Heat of Transformation

In Figure 12.21b, the phase changes appeared as horizontal line segments on the graph. During these segments, heat is being transferred to the system but the temperature isn't changing. The thermal energy continues to increase during a phase change, but, as noted, the additional energy goes into breaking molecular bonds rather than speeding up the molecules. **A phase change is characterized by a change in thermal energy without a change in temperature.**

The amount of heat energy that causes 1 kg of a substance to undergo a phase change is called the **heat of transformation** of that substance. For example, laboratory experiments show that 333,000 J of heat are needed to melt 1 kg of ice at 0°C. The symbol for heat of transformation is L . The heat required for the entire system of mass M to undergo a phase change is

$$Q = ML \quad (12.22)$$

“Heat of transformation” is a generic term that refers to any phase change. Two particular heats of transformation are the **heat of fusion** L_f , the heat of transformation between a solid and a liquid, and the **heat of vaporization** L_v , the heat of transformation between a liquid and a gas. The heat needed for these phase changes is

$$Q = \begin{cases} \pm ML_f & \text{Heat needed to melt/freeze mass } M \\ \pm ML_v & \text{Heat needed to boil/condense mass } M \end{cases} \quad (12.23)$$

where the \pm indicates that heat must be *added* to the system during melting or boiling but *removed* from the system during freezing or condensing. **You must explicitly include the minus sign when it is needed.**



Frozen frogs **BIO** It seems impossible, but common wood frogs survive the winter with much of their bodies frozen. When you dissolve substances in water, the freezing point lowers. Although the liquid water *between* cells in the frogs' bodies freezes, the water *inside* their cells remains liquid because of high concentrations of dissolved glucose. This prevents the cell damage that accompanies the freezing and subsequent thawing of tissues. When spring arrives, the frogs thaw and appear no worse for their winter freeze.



Lava—molten rock—undergoes a phase change from liquid to solid when it contacts liquid water; the transfer of heat to the water causes the water to undergo a phase change from liquid to gas.

Table 12.5 lists some heats of transformation. Notice that the heat of vaporization is always much higher than the heat of fusion, which our atomic model can explain. Melting breaks just enough molecular bonds to allow the system to lose rigidity and flow. Even so, the molecules in a liquid remain close together and loosely bonded. Vaporization breaks all bonds completely and sends the molecules flying apart. This process requires a larger increase in the thermal energy and thus a larger quantity of heat.

TABLE 12.5 Melting and boiling temperatures and heats of transformation at standard atmospheric pressure

Substance	T_m (°C)	L_f (J/kg)	T_b (°C)	L_v (J/kg)
Nitrogen (N_2)	-210	0.26×10^5	-196	1.99×10^5
Ethyl alcohol	-114	1.09×10^5	78	8.79×10^5
Mercury	-39	0.11×10^5	357	2.96×10^5
Water	0	3.33×10^5	100	22.6×10^5
Lead	328	0.25×10^5	1750	8.58×10^5

EXAMPLE 12.15 Melting a popsicle

A girl eats a 45 g frozen popsicle that was taken out of a -10°C freezer. How much energy does her body use to bring the popsicle up to body temperature?

PREPARE We can assume that the popsicle is pure water. Normal body temperature is 37°C. The specific heats of ice and liquid water are given in Table 12.4; the heat of fusion of water is given in Table 12.5.

SOLVE There are three parts to the problem: The popsicle must be warmed to 0°C, then melted, and then the resulting water must be warmed to body temperature. The heat needed to warm the frozen water by $\Delta T = 10^\circ\text{C} = 10\text{ K}$ to the melting point is

$$Q_1 = Mc_{\text{ice}} \Delta T = (0.045\text{ kg})(2090\text{ J/kg} \cdot \text{K})(10\text{ K}) = 940\text{ J}$$

Note that we use the specific heat of water ice, not liquid water, in this equation. Melting 45 g of ice requires heat

$$Q_2 = ML = (0.045\text{ kg})(3.33 \times 10^5\text{ J/kg}) = 15,000\text{ J}$$

The liquid water must now be warmed to body temperature; this requires heat

$$Q_3 = Mc_{\text{water}} \Delta T = (0.045\text{ kg})(4190\text{ J/kg} \cdot \text{K})(37\text{ K}) \\ = 7000\text{ J}$$

The total energy is the sum of these three values: $Q_{\text{total}} = 23,000\text{ J}$.

ASSESS More energy is needed to melt the ice than to warm the water, as we would expect. A commercial popsicle has 40 Calories, which is about 170 kJ. Roughly 15% of the chemical energy in this frozen treat is used to bring it up to body temperature!

Evaporation

Water boils at 100°C. But individual molecules of water can move from the liquid phase to the gas phase at lower temperatures. This process is known as **evaporation**. Water evaporates as sweat from your skin at a temperature well below 100°C. We can use our atomic model to explain this.

In gases, we have seen that there is a variation in particle speeds. The same principle applies to water and other liquids. At any temperature, some molecules will be moving fast enough to go into the gas phase. And they will do so, carrying away thermal energy as they go. The molecules that leave the liquid are the ones that have the highest kinetic energy, so evaporation reduces the average kinetic energy (and thus the temperature) of the liquid left behind.

The evaporation of water, both from sweat and in moisture you exhale, is one of the body's methods of exhausting the excess heat of metabolism to the environment, allowing you to maintain a steady body temperature. The heat to evaporate a mass M of water is $Q = ML_v$, and this amount of heat is removed from your body. However, the heat of vaporization L_v is a little larger than the value for boiling. At a skin



Keeping your cool **BIO** Humans (and cattle and horses) have sweat glands, so we can perspire to moisten our skin, allowing evaporation to cool our bodies. Animals that do not perspire can also use evaporation to keep cool. Dogs, goats, rabbits, and even birds pant, evaporating water from their respiratory passages. Elephants spray water on their skin; other animals may lick their fur.

temperature of 30°C , the heat of vaporization of water is $L_v = 24 \times 10^5 \text{ J/kg}$, or 6% higher than the value at 100°C in Table 12.5.

EXAMPLE 12.16 Computing heat loss by perspiration

The human body can produce approximately 30 g of perspiration per minute. At what rate is it possible to exhaust heat by the evaporation of perspiration?

SOLVE The evaporation of 30 g of perspiration at normal body temperature requires heat energy

$$Q = ML_v = (0.030 \text{ kg})(24 \times 10^5 \text{ J/kg}) = 7.2 \times 10^4 \text{ J}$$

This is the heat lost per minute; the rate of heat loss is

$$\frac{Q}{\Delta t} = \frac{7.2 \times 10^4 \text{ J}}{60 \text{ s}} = 1200 \text{ W}$$

ASSESS Given the metabolic power required for different activities, as listed in Chapter 11, this rate of heat removal is sufficient to keep the body cool even when exercising in hot weather—as long as the person drinks enough water to keep up this rate of perspiration.

STOP TO THINK 12.5 1 kg of barely molten lead, at 328°C , is poured into a large beaker holding liquid water right at the boiling point, 100°C . What is the mass of the water that will be boiled away as the lead solidifies?

- A. 0 kg B. $<1 \text{ kg}$ C. 1 kg D. $>1 \text{ kg}$

12.6 Calorimetry

At one time or another you've probably put an ice cube into a hot drink to cool it quickly. You were engaged, in a trial-and-error way, in a practical aspect of heat transfer known as **calorimetry**, the quantitative measurement of the heat transferred between systems or evolved in reactions. You know that heat energy will be transferred from the hot drink into the cold ice cube, reducing the temperature of the drink, as shown in **FIGURE 12.22**.

Let's make this qualitative picture more precise. **FIGURE 12.23** shows two systems that can exchange heat with each other but that are isolated from everything else. Suppose they start at different temperatures T_1 and T_2 . As you know, heat energy will be transferred from the hotter to the colder system until they reach a common final temperature T_f . In the cooling coffee example of Figure 12.22, the coffee is system 1, the ice is system 2, and the insulating barrier is the mug.

The insulation prevents any heat energy from being transferred to or from the environment, so energy conservation tells us that any energy leaving the hotter system must enter the colder system. The concept is straightforward, but to state the idea mathematically we need to be careful with signs.

Let Q_1 be the energy transferred to system 1 as heat. Q_1 is positive if energy *enters* system 1, negative if energy *leaves* system 1. Similarly, Q_2 is the energy transferred to system 2. The fact that the systems are merely exchanging energy can be written as $|Q_1| = |Q_2|$. But Q_1 and Q_2 have opposite signs, so $Q_1 = -Q_2$. No energy is exchanged with the environment, so it makes more sense to write this relationship as

$$Q_{\text{net}} = Q_1 + Q_2 = 0 \quad (12.24)$$

This idea is easily extended. No matter how many systems interact thermally, Q_{net} must equal zero.

NOTE ► The signs are very important in calorimetry problems. ΔT is always $T_f - T_i$, so ΔT and Q are negative for any system whose temperature decreases. The proper sign of Q for any phase change must be supplied by *you*, depending on the direction of the phase change. ◀

FIGURE 12.22 Cooling hot coffee with ice.

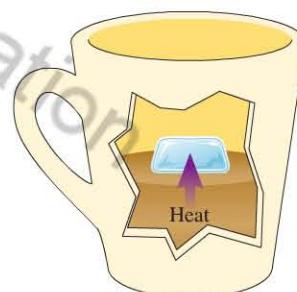


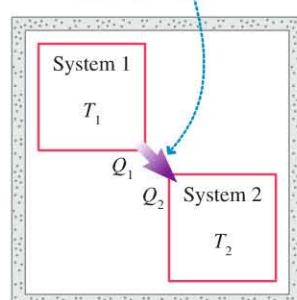
FIGURE 12.23 Two systems interact thermally. In this particular example, where heat is being transferred from system 1 to system 2, $Q_1 > 0$ and $Q_2 < 0$.

Heat energy is transferred from system 1 to system 2.
Energy conservation requires

$$|Q_1| = |Q_2|$$

Opposite signs mean that

$$Q_{\text{net}} = Q_1 + Q_2 = 0$$



PROBLEM-SOLVING STRATEGY 12.1

Calorimetry problems



PREPARE Identify the individual interacting systems. Assume that they are isolated from the larger environment. List known information and identify what you need to find. Convert all quantities to SI units.

SOLVE The statement of energy conservation is

$$Q_{\text{net}} = Q_1 + Q_2 + \dots = 0$$

- For systems that undergo a temperature change, $Q_{\Delta T} = Mc(T_f - T_i)$. Be sure to have the temperatures T_i and T_f in the correct order.
- For systems that undergo a phase change, $Q_{\text{phase}} = \pm ML$. Supply the correct sign by observing whether energy enters or leaves the system during the transition.
- Some systems may undergo a temperature change *and* a phase change. Treat the changes separately. The heat energy is $Q = Q_{\Delta T} + Q_{\text{phase}}$.

ASSESS The final temperature should be in the middle of the initial temperatures. A T_f that is higher or lower than all initial temperatures is an indication that something is wrong, usually a sign error.

EXAMPLE 12.17 Using calorimetry to identify a metal

200 g of an unknown metal is heated to 90.0°C, then dropped into 50.0 g of water at 20.0°C in an insulated container. The water temperature rises within a few seconds to 27.7°C, then changes no further. Identify the metal.

PREPARE The metal and the water interact thermally; there are no phase changes. We know all the initial and final temperatures. We will label the temperatures as follows: The initial temperature of the metal is T_m ; the initial temperature of the water is T_w . The common final temperature is T_f . For water, $c_w = 4190 \text{ J/kg} \cdot \text{K}$ is known from Table 12.4. Only the specific heat c_m of the metal is unknown.

SOLVE Energy conservation requires $Q_w + Q_m = 0$. Using $Q = Mc(T_f - T_i)$ for each, we have

$$Q_w + Q_m = M_w c_w (T_f - T_w) + M_m c_m (T_f - T_m) = 0$$

This is solved for the unknown specific heat:

$$\begin{aligned} c_m &= \frac{-M_w c_w (T_f - T_w)}{M_m (T_f - T_m)} \\ &= \frac{-(0.0500 \text{ kg})(4190 \text{ J/kg} \cdot \text{K})(27.7^\circ\text{C} - 20.0^\circ\text{C})}{(0.200 \text{ kg})(27.7^\circ\text{C} - 90.0^\circ\text{C})} \\ &= 129 \text{ J/kg} \cdot \text{K} \end{aligned}$$

Referring to Table 12.4, we find we have either 200 g of gold or, if we made an ever-so-slight experimental error, 200 g of lead!

ASSESS The temperature of the unknown metal changed much more than the temperature of the water. This means that the specific heat of the metal must be much less than that of water, which is exactly what we found.

EXAMPLE 12.18 Calorimetry with a phase change

Your 500 mL diet soda, with a mass of 500 g, is at 20°C, room temperature, so you add 100 g of ice from the -20°C freezer. Does all the ice melt? If so, what is the final temperature? If not, what fraction of the ice melts? Assume that you have a well-insulated cup.

PREPARE We need to distinguish between two possible outcomes. If all the ice melts, then $T_f > 0^\circ\text{C}$. It's also possible that the soda will cool to 0°C before all the ice has melted, leaving the ice and liquid in equilibrium at 0°C. We need to distinguish between these before knowing how to proceed. All the initial temperatures, masses, and specific heats are known. The final temperature of the combined soda + ice system is unknown.

SOLVE Let's first calculate the heat needed to melt all the ice and leave it as liquid water at 0°C. To do so, we must warm the ice by 20 K to 0°C, then change it to water. The heat input for this two-stage process is

This is the energy to raise the temperature of the ice from -20°C to 0°C. $\Delta T = 20 \text{ K}$.

This is the energy to melt the ice once it reaches 0°C.

$$Q_{\text{melt}} = \underbrace{M_{\text{ice}} c_{\text{ice}} (20 \text{ K})}_{\text{heat to raise temp}} + \underbrace{M_{\text{ice}} L_f}_{\text{heat to melt}} = 37,000 \text{ J}$$

where L_f is the heat of fusion of water.

Q_{melt} is a *positive* quantity because we must *add* heat to melt the ice. Next, let's calculate how much heat energy will leave the 500 g soda if it cools all the way to 0°C:

$$Q_{\text{cool}} = M_{\text{soda}} c_{\text{water}} (-20 \text{ K}) = -42,000 \text{ J}$$

where $\Delta T = -20 \text{ K}$ because the temperature decreases. Because $|Q_{\text{cool}}| > Q_{\text{melt}}$, the soda has sufficient energy to melt all the ice. Hence the final state will be all liquid at $T_f > 0$. (Had we found

$|Q_{\text{cool}}| < Q_{\text{melt}}$, then the final state would have been an ice-liquid mixture at 0°C .)

Energy conservation requires $Q_{\text{ice}} + Q_{\text{soda}} = 0$. The heat Q_{ice} consists of three terms: warming the ice to 0°C , melting the ice to water at 0°C , then warming the 0°C water to T_f . The mass will still be M_{ice} in the last of these steps because it is the “ice system,” but we need to use the specific heat of *liquid water*. Thus

$$\begin{aligned} Q_{\text{ice}} + Q_{\text{soda}} &= [M_{\text{ice}} c_{\text{ice}}(20 \text{ K}) + M_{\text{ice}} L_f \\ &\quad + M_{\text{ice}} c_{\text{water}}(T_f - 0^{\circ}\text{C})] \\ &\quad + M_{\text{soda}} c_{\text{water}}(T_f - 20^{\circ}\text{C}) = 0 \end{aligned}$$

We've already done part of the calculation, allowing us to write

$$37,000 \text{ J} + M_{\text{ice}} c_{\text{water}}(T_f - 0^{\circ}\text{C}) + M_{\text{soda}} c_{\text{water}}(T_f - 20^{\circ}\text{C}) = 0$$

Solving for T_f gives

$$T_f = \frac{20M_{\text{soda}}c_{\text{water}} - 37,000}{M_{\text{ice}}c_{\text{water}} + M_{\text{soda}}c_{\text{water}}} = 1.9^{\circ}\text{C}$$

ASSESS A good deal of ice has been put in the soda, so it ends up being cooled nearly to the freezing point, as we might expect.

STOP TO THINK 12.6 1 kg of lead at 100°C is dropped into a container holding 1 kg of water at 0°C . Once the lead and water reach thermal equilibrium, the final temperature is

- A. $<50^{\circ}\text{C}$ B. 50°C C. $>50^{\circ}\text{C}$

12.7 Thermal Properties of Gases

Now it's time to turn back to gases. What happens when they are heated? Can we assign specific heats to gases as we did for solids and liquids? As we'll see, gases are harder to characterize than solids or liquids because the heat required to cause a specified temperature change depends on the *process* by which the gas changes state.

Just as for solids and liquids, heating a gas changes its temperature. But by how much? **FIGURE 12.24** shows two isotherms on the pV diagram for a gas. Processes A and B, which start on the T_i isotherm and end on the T_f isotherm, have the *same* temperature change $\Delta T = T_f - T_i$, so we might expect them both to require the same amount of heat. But it turns out that process A, which takes place at constant volume, requires *less* heat than does process B, which occurs at constant pressure. The reason is that work is done in process B but not in process A.

It is useful to define two different versions of the specific heat of gases: one for constant-volume processes and one for constant-pressure processes. We will define these as *molar* specific heats because we usually do gas calculations using moles instead of mass. The quantity of heat needed to change the temperature of n moles of gas by ΔT is, for a constant-volume process,

$$Q = nC_V \Delta T \quad (12.25)$$

and for a constant-pressure process,

$$Q = nC_P \Delta T \quad (12.26)$$

C_V is the **molar specific heat at constant volume**, and C_P is the **molar specific heat at constant pressure**. Table 12.6 on the next page gives the values of C_V and C_P for some common gases. The units are $\text{J/mol} \cdot \text{K}$. The values for air are essentially equal to those for N_2 .

It's interesting that all the monatomic gases in Table 12.6 have the same values for C_P and for C_V . Why should this be? Monatomic gases are really close to ideal, so let's go back to our atomic model of the ideal gas. We know that the thermal energy of an ideal gas of N atoms is $E_{\text{th}} = \frac{3}{2}Nk_B T = \frac{3}{2}nRT$. If the temperature of an ideal gas changes by ΔT , its thermal energy changes by

$$\Delta E_{\text{th}} = \frac{3}{2}nR \Delta T \quad (12.27)$$



8.7

FIGURE 12.24 Processes A and B have the same ΔT and the same ΔE_{th} , but they require different amounts of heat.

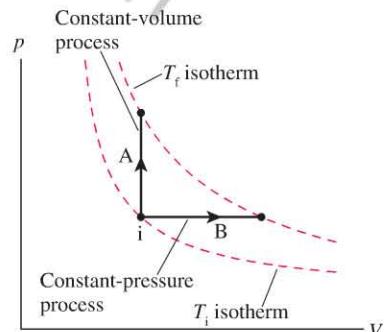


TABLE 12.6 Molar specific heats of gases (J/mol · K) at 20°C

Gas	C_P	C_V
Monatomic Gases		
He	20.8	12.5
Ne	20.8	12.5
Ar	20.8	12.5
Diatomeric Gases		
H ₂	28.7	20.4
N ₂	29.1	20.8
O ₂	29.2	20.9
Triatomic Gas		
Water vapor	33.3	25.0

If we keep the volume of the gas constant, so that no work is done, this energy change can come only from heat, so

$$Q = \frac{3}{2}nR \Delta T \quad (12.28)$$

Comparing Equation 12.28 with the definition of molar specific heat in Equation 12.25, we see that the molar specific heat at constant volume must be

$$C_V(\text{monatomic gas}) = \frac{3}{2}R = 12.5 \text{ J/mol} \cdot \text{K} \quad (12.29)$$

This *predicted* value from the ideal-gas model is exactly the *measured* value of C_V for the monatomic gases in Table 12.6, a good check that the model we have been using is correct.

The molar specific heat at constant pressure is different. If you heat a gas in a sealed container so that there is no change in volume, then no work is done. But if you heat a sample of gas in a cylinder with a piston to keep it at constant pressure, the gas must expand, and it will do work as it expands. The expression for ΔE_{th} in Equation 12.27 is still valid, but now, according to the first law of thermodynamics, $Q = \Delta E_{\text{th}} + W_{\text{gas}}$. The work done by the gas in an isobaric process is $W_{\text{gas}} = p \Delta V$, so the heat required is

$$Q = \Delta E_{\text{th}} + W_{\text{gas}} = \frac{3}{2}nR \Delta T + p \Delta V \quad (12.30)$$

The ideal-gas law, $pV = nRT$, implies that if p is constant and only V and T change, then $p \Delta V = nR \Delta T$. Using this result in Equation 12.30, we find the heat needed to change the temperature by ΔT in a constant-pressure process is

$$Q = \frac{3}{2}nR \Delta T + nR \Delta T = \frac{5}{2}nR \Delta T$$

A comparison with the definition of molar specific heat shows that

$$C_P = \frac{5}{2}R = 20.8 \text{ J/mol} \cdot \text{K} \quad (12.31)$$

This is larger than C_V , as expected, and in perfect agreement for all the monatomic gases in Table 12.6.

EXAMPLE 12.19 Work done by an expanding gas

A typical weather balloon is made of a thin latex envelope that takes relatively little force to stretch, so the pressure inside the balloon is approximately equal to atmospheric pressure. The balloon is filled with a gas that is less dense than air, typically hydrogen or helium. Suppose a weather balloon filled with 180 mol of helium is waiting for launch on a cold morning at a high-altitude station. The balloon warms in the sun, which raises the temperature of the gas from 0°C to 30°C. As the balloon expands, how much work is done by the expanding gas?

PREPARE The work done is equal to $p\Delta V$, but we don't know the pressure (it's not sea level and we don't know the altitude) and we don't know the volume of the balloon. Instead, we'll use the first law of thermodynamics. We can rewrite Equation 12.17 as

$$W_{\text{gas}} = Q - \Delta E_{\text{th}}$$

The change in temperature of the gas is 30°C, so $\Delta T = 30 \text{ K}$. We can compute how much heat energy is transferred to the balloon as it warms because this is a temperature change at constant

pressure, and we can compute how much the thermal energy of the gas increases because we know ΔT .

SOLVE The heat required to increase the temperature of the gas is given by Equation 12.26:

$$Q = nC_P \Delta T = (180 \text{ mol})(20.8 \text{ J/mol} \cdot \text{K})(30 \text{ K}) = 110 \text{ kJ}$$

The change in thermal energy depends on the change in temperature according to Equation 12.18:

$$\begin{aligned} E_{\text{th}} &= \frac{3}{2}nR \Delta T = \frac{3}{2}(180 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})(30 \text{ K}) \\ &= 67 \text{ kJ} \end{aligned}$$

The work done by the expanding balloon is just the difference between these two values:

$$W_{\text{gas}} = Q - \Delta E_{\text{th}} = 110 \text{ kJ} - 67 \text{ kJ} = 43 \text{ kJ}$$

ASSESS The numbers are large—it's a lot of heat and a large change in thermal energy—but it's a big balloon with a lot of gas, so this seems reasonable.

Now let's turn to the diatomic gases in Table 12.6. The molar specific heats are higher than for monatomic gases, and our atomic model explains why. The thermal energy of a monatomic gas consists exclusively of the translational kinetic energy of the atoms; heating a monatomic gas simply means that the atoms move faster. The thermal energy of a diatomic gas is more than just the translational energy, as shown in **FIGURE 12.25**. Heating a diatomic gas makes the molecules move faster, but it also causes them to rotate more rapidly. Energy goes into the translational kinetic energy of the molecules (thus increasing the temperature), but some goes into rotational kinetic energy. Because some of the added heat goes into rotation and not into translation, the specific heat of a diatomic gas is higher than that of a monatomic gas, as we see in Table 12.6. The table shows that water vapor, a triatomic gas, has an even higher specific heat, as we might expect.

Heat Engines

Chapter 11 introduced the idea of a *heat engine*, a device that takes in heat energy from a hot reservoir, does useful work, and exhausts waste heat to a cold reservoir. Many real heat engines rely on gas processes, and we've now reached the point where we can examine heat engines more closely. A diagram of the (somewhat simplified) operation of a real heat engine, a two-stroke gasoline engine, is shown in **FIGURE 12.26**.

FIGURE 12.26 A simplified model of a gasoline engine.

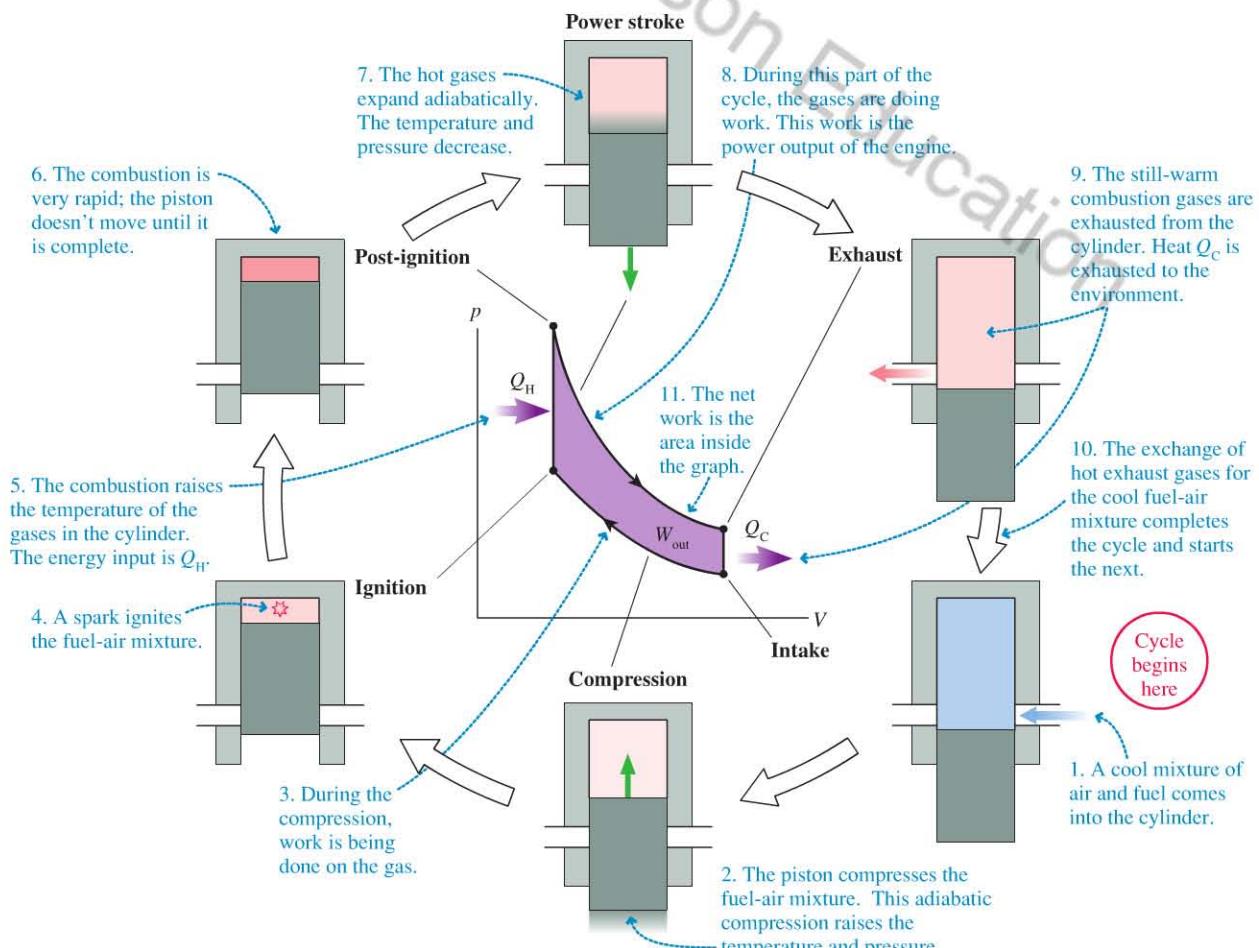
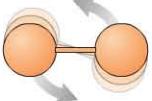
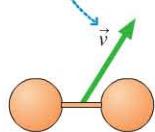


FIGURE 12.25 Thermal energy of a diatomic gas.

The thermal energy of a diatomic gas is the sum of the translational kinetic energy of the molecules ...

... plus the rotational kinetic energy of the molecules.



NOTE ► It is common to use “gas” as shorthand for “gasoline.” In this chapter, “gas” always refers to the state of matter. We will use “gasoline” or “fuel” to refer to the chemical energy source that you put in your tank. ◀

In this engine, the gases in the cylinder expand and contract in a repeating cycle as the piston moves back and forth. The cycle begins at the noted point with the introduction of gases (a mixture of air and fuel) into the cylinder. From this point, we can use our knowledge of gas processes to analyze the operation of the engine as it goes through its cycle. Figure 12.26 shows the connection between the operation of the engine and relevant points on the pV diagram of the gases in the engine's cylinder.

There are several important characteristics of this heat engine that apply generally to all heat engines:

- The sequence of processes forms a closed loop in the pV diagram. This must be true for the engine to go through its cycle over and over.
- Heat Q_H is taken in from a hot reservoir (in this case, from the burning fuel) at some point during the cycle. To bring the temperature back down at the end of the cycle, heat Q_C must be exhausted to a cold reservoir. This heat is carried out of the engine by the hot exhaust gases.
- The work done by the engine, W_{out} , is the area *inside* the closed loop on the pV diagram.
- Because the temperature returns to its starting point at the end of each cycle, there's no net change in thermal energy. Thus from energy conservation, $W_{\text{out}} = Q_H - Q_C$.

Work is the area inside the closed loop because of our earlier observation that the work done in a single process is the area under the graph. The work is negative for the compression, because ΔV is negative, and positive for the expansion. Thus the net work is the area under the upper curve *minus* the area under the lower curve, which is just the area inside the closed loop. This is true for all heat engines, not just the example shown here.

8.12, 8.13, 8.14 

EXAMPLE 12.20 Finding the pressure and temperature in a gasoline engine

At the start of a cycle in an internal combustion engine, 0.021 mol of air is introduced into the cylinder along with 0.018 g of gasoline. The adiabatic compression stroke raises the pressure to 17 atm and the temperature to 300°C. A spark then ignites the gasoline, which burns fully. What are the pressure and temperature immediately after ignition? (The combustion consumes oxygen but produces other gases; assume that the total number of moles of gas doesn't change.)

PREPARE Chapter 11 gives the energy content of gasoline as 44 kJ per gram, so we can figure out how much energy is liberated when the gas in the cylinder is burned. This combustion happens rapidly, so there is no time for the piston to move. The gas is heated at constant volume. The gas in the cylinder is mostly air, so we'll use $C_V = 20.8 \text{ J/mol} \cdot \text{K}$ as the molar specific heat at constant volume.

SOLVE The energy released in the combustion of 0.018 g of gasoline is

$$Q = (44 \text{ kJ/g})(0.018 \text{ g}) = 790 \text{ J}$$

This heat is added to the air in the cylinder at constant volume, causing the temperature to increase by

$$\Delta T = \frac{Q}{nC_V} = \frac{790 \text{ J}}{(0.021 \text{ mol})(20.8 \text{ J/mol} \cdot \text{K})} = 1800^\circ\text{C}$$

Thus igniting the fuel raises the temperature from $T_i = 300^\circ\text{C}$ at the end of the compression stroke to $T_f = 2100^\circ\text{C}$. The volume doesn't change, so the new pressure is

$$p_f = p_i \frac{V_i}{V_f} \frac{T_f}{T_i} = (17 \text{ atm})(1) \frac{2100 + 273}{300 + 273} = 70 \text{ atm}$$

ASSESS The pressure and temperature are both very large, but that's to be expected following an explosion in a confined volume. When you drive your car, a series of such explosions in your engine's cylinders provide the necessary force to move your car down the road. In a typical car moving at highway speeds, over 100 explosions are taking place under your hood each second.

12.8 Heat Transfer

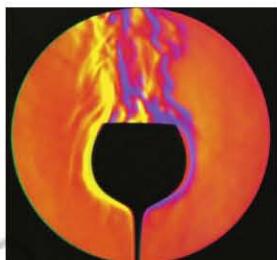
You feel warmer when the sun is shining on you, colder when you are sitting on a cold concrete bench or when a stiff wind is blowing. This is due to the transfer of heat. Although we've talked about heat and heat transfers a lot in these last two chapters, we haven't said much about *how* heat is transferred from a hotter object to a colder object. There are four basic mechanisms, described in the table on the next page,

by which objects exchange heat with other objects or their surroundings. Evaporation was treated in an earlier section; in this section we will consider the other mechanisms.

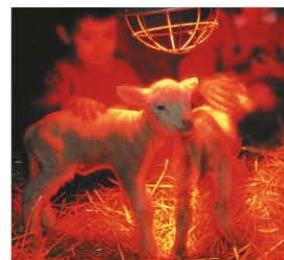
Heat-transfer mechanisms



When two objects are in direct physical contact, such as the soldering iron and the circuit board, heat is transferred by **conduction**. **Energy is transferred by direct contact.**



This special photograph shows air currents near a warm glass of water. Air near the glass is warmed and rises, taking thermal energy with it in a process known as **convection**. **Energy is transferred by the bulk motion of molecules with high thermal energy.**



The lamp shines on the lambs huddled below, warming them. The energy is transferred by infrared **radiation**, a form of electromagnetic waves. **Energy is transferred by electromagnetic waves.**



As we saw in an earlier section, the **evaporation** of liquid can carry away significant quantities of thermal energy. When you blow on a cup of cocoa, this increases the rate of evaporation, rapidly cooling it. **Energy is transferred by the removal of molecules with high thermal energy.**

Conduction

If you hold a metal spoon in a cup of hot coffee, the handle of the spoon soon gets warm. Thermal energy is transferred along the spoon from the coffee to your hand. The difference in temperature between the two ends drives this heat transfer by a process known as **conduction**. Conduction is the transfer of thermal energy directly through a physical material.

FIGURE 12.27 shows a copper rod placed between a hot reservoir (a fire) and a cold reservoir (a block of ice). We can use our atomic model to see how thermal energy is transferred along the rod by the interaction between atoms in the rod; fast-moving atoms at the hot end transfer energy to slower-moving atoms at the cold end.

Suppose we set up a series of experiments to measure the heat Q transferred through various rods. We would find the following trends in our data:

- Q increases if the temperature difference ΔT between the hot end and the cold end is increased.
- Q increases if the cross-section area A of the rod is increased.
- Q decreases if the length L of the rod is increased.
- Some materials (such as metals) transfer heat quite readily. Other materials (such as wood) transfer very little heat.

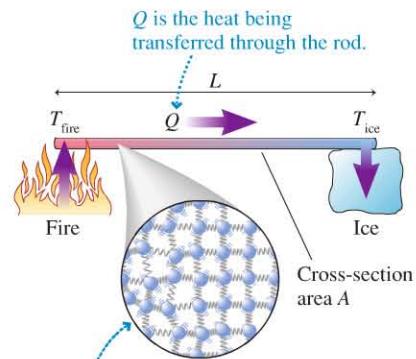
The final observation is one that is familiar to you: If you are stirring a pot of hot soup on the stove, you generally use a wood or plastic spoon rather than a metal one.

These experimental observations about heat conduction can be summarized in a single formula. If heat Q is transferred in a time interval Δt , the *rate* of heat transfer (joules per second) is $Q/\Delta t$. For a material of cross-section area A and length L , spanning a temperature difference ΔT , the rate of heat transfer is

$$\frac{Q}{\Delta t} = \left(\frac{kA}{L} \right) \Delta T \quad (12.32)$$

Rate of conduction of heat across a temperature difference

FIGURE 12.27 Conduction of heat in a solid rod.



The particles on the left side of the rod are vibrating more vigorously than the particles on the right. The particles on the left transfer energy to the particles on the right via the bonds connecting them.

TABLE 12.7 Thermal conductivity values (measured at 20°C)

Material	<i>k</i> (W/m · K)
Diamond	1000
Silver	420
Copper	400
Iron	72
Stainless steel	14
Ice	1.7
Concrete	0.8
Plate glass	0.75
Skin	0.50
Muscle	0.46
Fat	0.21
Wood	0.2
Carpet	0.04
Fur, feathers	0.02–0.06
Air (27°C, 100 kPa)	0.026



Carpet or tile? If you walk around your house barefoot, you will notice that tile floors feel much colder than carpeted floors. The tile floor isn't really colder—the temperatures of all the floors in your house are nearly the same. But tile has a much higher thermal conductivity than carpet, so more heat will flow from your feet into the tile than into carpet. This causes the tile to feel colder.

The quantity *k*, which characterizes whether the material is a good conductor of heat or a poor conductor, is called the **thermal conductivity** of the material. Because the heat transfer rate J/s is a *power*, measured in watts, the units of *k* are W/m · K. Values of *k* for some common materials are listed in Table 12.7; a larger number for *k* means a material is a better conductor of heat.

Good conductors of electricity, such as silver and copper, are usually good conductors of heat. Air, like all gases, is a poor conductor of heat because there are no bonds between adjacent atoms.

The weak bonds between molecules make most biological materials poor conductors of heat. Muscle is a better conductor than fat, so sea mammals have thick layers of fat for insulation. Land mammals insulate their bodies with fur, birds with feathers. Both fur and feathers trap a good deal of air, so their conductivity is similar to that of air, as Table 12.7 shows.

EXAMPLE 12.21 Heat loss by conduction

At the start of the section we noted that you “feel cold” when you sit on a cold concrete bench. “Feeling cold” really means that your body is losing a significant amount of heat. How significant? Suppose you are sitting on a 10°C concrete bench. You are wearing thin clothing that provides negligible insulation. In this case, most of the insulation that protects your body’s core (temperature 37°C) from the cold of the bench is provided by a 1.0-cm-thick layer of fat on the part of your body that touches the bench. (The thickness varies from person to person, but this is a reasonable average value.) A good estimate of the area of contact with the bench is 0.10 m². Given these details, what is the heat loss by conduction?

PREPARE Heat is lost to the bench by conduction through the fat layer, so we can compute the rate of heat loss by using Equation 12.32. The thickness of the conducting layer is 0.010 m, the area is 0.10 m², and the thermal conductivity of fat is given in Table 12.7. The temperature difference is the difference between your body’s core temperature (37°C) and the temperature of the bench (10°C), a difference of 27°C, or 27 K.

SOLVE We have all of the data we need to use Equation 12.32 to compute the rate of heat loss:

$$\frac{Q}{\Delta t} = \left(\frac{(0.21 \text{ W/m} \cdot \text{K})(0.10 \text{ m}^2)}{0.010 \text{ m}} \right)(27 \text{ K}) = 57 \text{ W}$$

ASSESS 57 W is more than half your body’s resting power, which we learned in Chapter 11 is approximately 100 W. That’s a significant loss, so your body will feel cold, a result that seems reasonable if you’ve ever sat on a cold bench for any length of time.

Convection

We noted that air is a poor conductor of heat. In fact, the data in Table 12.7 show that air has a thermal conductivity comparable to that of feathers. So why, on a cold day, will you be more comfortable if you are wearing a down jacket?

In conduction, faster-moving atoms transfer thermal energy to adjacent atoms. But in fluids such as water or air, there is a more efficient means to move energy: by transferring the faster-moving atoms themselves. When you place a pan of cold water on a burner on the stove, it’s heated on the bottom. This heated water expands and becomes less dense than the water above it, so it rises to the surface while cooler, denser water sinks to take its place. This transfer of thermal energy by the motion of a fluid is known as **convection**.

Convection is usually the main mechanism for heat transfer in fluid systems. On a small scale, convection mixes the pan of water that you heat on the stove; on a large



Warm water (colored) moves by convection.

scale, convection is responsible for making the wind blow and ocean currents circulate. Air is a very poor thermal conductor, but it is very effective at transferring energy by convection. To use air for thermal insulation, it is necessary to trap the air in small pockets to limit convection. And that's exactly what feathers, fur, double-paned windows, and fiberglass insulation do. Convection is much more rapid in water than in air, which is why people can die of hypothermia in 65°F water but can live quite happily in 65°F air.

Radiation

You *feel* the warmth from the glowing red coals in a fireplace. On a cool day, you prefer to sit in the sun rather than the shade so that the sunlight keeps you warm. In both cases heat energy is being transferred to your body in the form of **radiation**.

Radiation consists of electromagnetic waves—a topic we will explore further in later chapters—that transfer energy from the object that emits the radiation to the object that absorbs it. Visible light, the red-to-violet colors of the rainbow, consists of electromagnetic waves, and objects that are hot enough to glow emit part of their radiation as visible light. We say these objects are “red hot” or, at a high enough temperature, “white hot.” The white light from an incandescent lightbulb is radiation emitted by a thin wire filament that has been heated to a very high temperature by an electric current.

Electromagnetic waves of somewhat lower frequency are called *infrared radiation*; we can’t see these waves, but we can sometimes sense them as “heat” as they are absorbed on our skin. Glowing objects emit part of their radiation as visible light but much of it as infrared radiation—the warmth that you feel from the glowing coals. Objects near room temperature, including your body, also radiate, but they emit nearly all of their radiation as infrared radiation. Thermal images, like the one of the teapot in FIGURE 12.28, are made with cameras having special infrared-sensitive detectors that allow them to “see” radiation that our eyes don’t respond to.

NOTE ► The word “radiation” comes from “radiate,” which means “to beam.” You have likely heard the word used to refer to x rays and radioactive materials. This is not the sense we will use in this chapter. Here, we will use “radiation” to mean electromagnetic waves that “beam” from an object. ◀

Radiation is a significant part of the energy balance that keeps your body at the proper temperature. Radiation from the sun or from warm, nearby objects is absorbed by your skin, increasing your body’s thermal energy. At the same time, your body radiates energy away. You can change your body temperature by absorbing more radiation (sitting next to a fire) or by emitting more radiation (taking off your hat and scarf to expose more skin).

The energy radiated by an object shows a strong dependence on temperature. In Figure 12.28, the hot teapot glows much more brightly than the cooler hand. In the photo that opened the chapter, the person’s warm hands emit quite a bit more energy than the cooler spider. We can quantify the dependence of radiated energy on temperature. If heat energy Q is radiated in a time interval Δt by an object with surface area A and absolute temperature T , the *rate* of heat transfer $Q/\Delta t$ (joules per second) is found to be

$$\frac{Q}{\Delta t} = e\sigma AT^4 \quad (12.33)$$

Rate of heat transfer by radiation at temperature T (Stefan’s law)

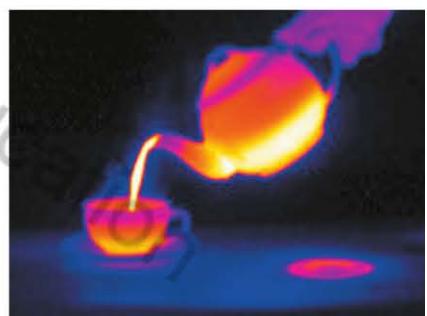
Quantities in this equation are defined as follows:

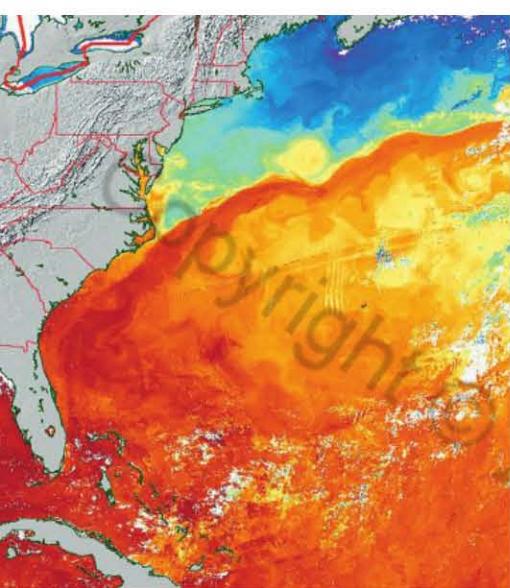
- e is the **emissivity** of a surface, a measure of the effectiveness of radiation. The value of e ranges from 0 to 1. Human skin is a very effective radiator at body temperature, with $e = 0.97$, regardless of skin color.



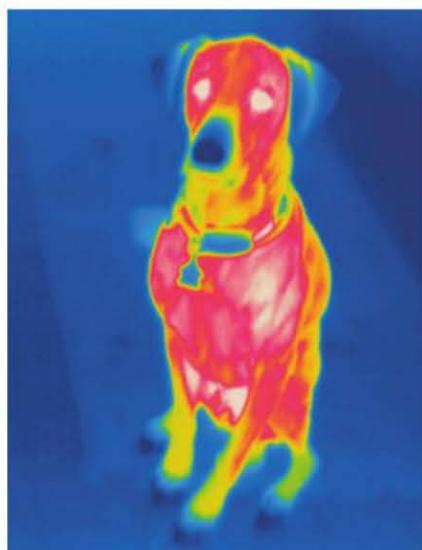
A feather coat BIO Penguins are birds and thus have feathers, even though they do not fly. The feathers are more obvious on this penguin chick than on the adults. A penguin’s short, dense feathers serve a different role than the flight feathers of other birds: They trap air to provide thermal insulation. The feathers are equipped with muscles at the base of the shafts to flatten them and eliminate these air pockets; otherwise, the penguins would be too buoyant to swim underwater.

FIGURE 12.28 A thermal image of a teapot.





Global heat transfer This satellite image shows radiation emitted by the waters of the ocean off the east coast of the United States. You can clearly see the warm waters of the Gulf Stream, a large-scale convection that transfers heat to northern latitudes. The satellite “sees” the radiation from the earth, so this radiation must readily pass through the atmosphere into space. In fact, the *only* means by which the earth as a whole can reduce its thermal energy is by radiation, because there is no conduction or convection in the vacuum of space. This has consequences for the energy balance of the earth, as we consider in a section on the greenhouse effect at the end of Part III.



The infrared radiation emitted by a dog is captured in this image. His cool nose and paws radiate much less energy than the rest of his body.

- T is the absolute temperature in kelvin.
- A is the surface area in m^2 .
- σ is a constant known as the Stefan-Boltzmann constant, with the value $\sigma = 5.67 \times 10^{-8} \frac{\text{W}}{\text{m}^2 \cdot \text{K}^4}$.

Notice the very strong fourth-power dependence on temperature. Doubling the absolute temperature of an object increases the radiant heat transfer by a factor of 16!

The amount of energy radiated by an animal can be surprisingly large. An adult human with bare skin in a room at a comfortable temperature has a skin temperature of approximately 33°C , or 306 K. A typical value for the skin’s surface area is 1.8 m^2 . With these values and the emissivity of skin noted above, we can calculate the rate of heat transfer via radiation from the skin:

$$\frac{Q}{\Delta t} = e\sigma AT^4 = (0.97) \left(5.67 \times 10^{-8} \frac{\text{W}}{\text{m}^2 \cdot \text{K}^4} \right) (1.8 \text{ m}^2) (306 \text{ K})^4 = 870 \text{ W}$$

As we learned in Chapter 11, the body at rest generates approximately 100 W of thermal energy. If the body radiated energy at 870 W, it would quickly cool. At this rate of emission, the body temperature would drop by 1°C every 7 minutes! Clearly, there must be some mechanism to balance this emitted radiation, and there is: the radiation *absorbed* by the body.

When you sit in the sun, your skin warms due to the radiation you absorb. Even if you are not in the sun, you are absorbing the radiation emitted by the objects surrounding you. Suppose an object at temperature T is surrounded by an environment at temperature T_0 . The *net* rate at which the object radiates heat energy—that is, radiation emitted minus radiation absorbed—is

$$\frac{Q_{\text{net}}}{\Delta t} = e\sigma A(T^4 - T_0^4) \quad (12.34)$$

This makes sense. An object should have no net energy transfer by radiation if it’s in thermal equilibrium ($T = T_0$) with its surroundings. Note that the emissivity e appears for absorption as well; objects that are good emitters are also good absorbers.

EXAMPLE 12.22 Determining energy loss by radiation for the body

A person with a skin temperature of 33°C is in a room at 24°C . What is the net rate of heat transfer by radiation?

PREPARE Body temperature is $T = 33 + 273 = 306 \text{ K}$; the temperature of the room is $T_0 = 24 + 273 = 297 \text{ K}$.

SOLVE The net radiation rate, given by Equation 12.34, is

$$\begin{aligned} \frac{Q_{\text{net}}}{\Delta t} &= e\sigma A(T^4 - T_0^4) \\ &= (0.97) \left(5.67 \times 10^{-8} \frac{\text{W}}{\text{m}^2 \cdot \text{K}^4} \right) (1.8 \text{ m}^2) [(306 \text{ K})^4 - (297 \text{ K})^4] = 98 \text{ W} \end{aligned}$$

ASSESS This is a reasonable value, roughly matching your resting metabolic power. When you are dressed (little convection) and sitting on wood or plastic (little conduction), radiation is your body’s primary mechanism for dissipating the excess thermal energy of metabolism.

STOP TO THINK 12.7 Suppose you are an astronaut in the vacuum of space, hard at work in your sealed spacesuit. The only way that you can transfer heat to the environment is by

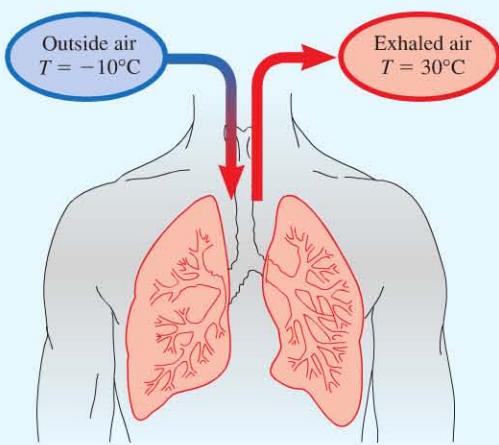
- A. Conduction. B. Convection. C. Radiation. D. Evaporation.

INTEGRATED EXAMPLE 12.23**Breathing in cold air**

On a cold day, breathing costs your body energy; as the cold air comes into contact with the warm tissues of your lungs, the air warms due to heat transferred from your body. The thermal image of the person inflating a balloon in **FIGURE 12.29** shows that exhaled air is quite a bit warmer than the surroundings.

FIGURE 12.30 shows this process for a frosty -10°C (14°F) day. The inhaled air warms to nearly the temperature of the interior of your body, 37°C . When you exhale, some heat is retained by the body, but most is lost; the exhaled air is still about 30°C .

FIGURE 12.30 Breathing warms the air.



Your lungs hold several liters of air, but only a small part of the air is exchanged during each breath. A typical person takes 12 breaths each minute, with each breath drawing in 0.50 L of outside air. If the air warms up from -10°C to 30°C , what is the volume of the air exhaled with each breath? What fraction of your body's resting power goes to warming the air? (Note that gases are exchanged as you breathe—oxygen to carbon dioxide—but to a good approximation the number of atoms, and thus the number of moles, stays the same.) Consider only the energy required to warm the air, not the energy lost to evaporation from the tissues of the lungs.

PREPARE There are two parts to the problem, which we'll solve in two different ways. First, we need to figure out the volume of exhaled air. When your body warms the air, its temperature

FIGURE 12.29 A thermal image of a person blowing up a balloon on a cold day.



increases, so its volume will increase as well. The initial and final states represented in Figure 12.30 are at atmospheric pressure, so we can treat this change as an isobaric (constant-pressure) process. We'll need the absolute temperatures of the initial and final states:

$$T_i = -10^{\circ}\text{C} + 273 = 263 \text{ K}$$

$$T_f = 30^{\circ}\text{C} + 273 = 303 \text{ K}$$

Next, we'll determine how much heat energy is required to raise the temperature of the air. Because the initial and final pressures are the same, we'll do this by computing the heat needed to raise the temperature of a gas at constant pressure. The change in temperature is $+40^{\circ}\text{C}$, so we can use $\Delta T = 40 \text{ K}$. Air is a mix of nitrogen and oxygen with a small amount of other gases. C_p for nitrogen and oxygen is the same to two significant figures, so we will assume the gas has $C_p = 29 \text{ J/mol} \cdot \text{K}$.

SOLVE The change in volume of the air in your lungs is a constant-pressure process. 0.50 L of air is breathed in; this is V_i . When the temperature increases, so does the volume. The gas isn't in a sealed container, but we are considering the same "parcel" of gas before and after, so we can use the ideal-gas law to find the volume after the temperature increase:

$$V_f = V_i \frac{p_f T_f}{p_i T_i} = (0.50 \text{ L}) \times 1 \times \frac{303 \text{ K}}{263 \text{ K}} = 0.58 \text{ L}$$

The volume increases by a little over 10%, from 0.50 L to 0.58 L.

Now we can move on to the second part: determining the energy required. We'll need the number of moles of gas, which we can compute from the ideal-gas law:

$$n = \frac{pV}{RT} = \frac{(101.3 \times 10^3 \text{ Pa})(0.50 \times 10^{-3} \text{ m}^3)}{(8.31 \text{ J/mol} \cdot \text{K})(263 \text{ K})} = 0.023 \text{ mol}$$

In doing this calculation, we used $1 \text{ m}^3 = 1000 \text{ L}$ to convert the 0.50 L volume to m^3 . Now we can compute the heat needed to warm one breath, using Equation 12.26:

$$\begin{aligned} Q(\text{one breath}) &= n C_p \Delta T = (0.023 \text{ mol})(29 \text{ J/mol} \cdot \text{K})(40 \text{ K}) \\ &= 27 \text{ J} \end{aligned}$$

If you take 12 breaths a minute, a single breath takes 1/12 of a minute, or 5.0 s. Thus the heat power your body supplies to warm the incoming air is

$$P = \frac{Q}{\Delta t} = \frac{27 \text{ J}}{5.0 \text{ s}} = 5.4 \text{ W}$$

At rest, your body typically uses 100 W, so this is just over 5% of your body's resting power. You breathe in a good deal of air each minute and warm it by quite a bit, but the specific heat of air is small enough that the energy required is reasonably modest.

ASSESS The air expands slightly and a small amount of energy goes into heating it. If you've been outside on a cold day, you know that neither change is dramatic, so this final result seems reasonable. The energy loss is noticeable but reasonably modest; other forms of energy loss are more important when you are outside on a cold day.

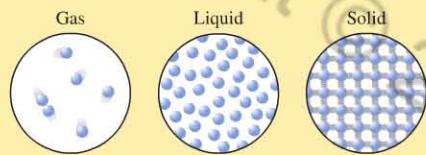
SUMMARY

The goal of Chapter 12 has been to use the atomic model of matter to explain and explore many macroscopic phenomena associated with heat, temperature, and the properties of matter.

GENERAL PRINCIPLES

Atomic Model

We model matter as being made of simple basic particles. The relationship of these particles to each other defines the phase.



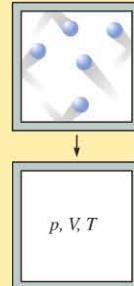
The atomic model explains thermal expansion, specific heat, and heat transfer.

Atomic Model of a Gas

Macroscopic properties of gases can be explained in terms of the atomic model of the gas. The speed of the particles is related to the temperature:

$$v_{\text{rms}} = \sqrt{\frac{3k_B T}{m}}$$

The collisions of particles with each other and with the walls of the container determine the pressure.



Ideal-Gas Law

The ideal gas law relates the pressure, volume, and temperature in a sample of gas. We can express the law in terms of the number of atoms or the number of moles in the sample:

$$pV = Nk_B T$$

$$pV = nRT$$

For a gas process in a sealed container,

$$\frac{p_i V_i}{T_i} = \frac{p_f V_f}{T_f}$$

IMPORTANT CONCEPTS

Effects of heat transfer

A system that is heated can either change temperature or change phase.

The **specific heat c** of a material is the heat required to raise 1 kg by 1 K.

$$Q = Mc \Delta T$$

The **heat of transformation** is the energy necessary to change the phase of 1 kg of a substance. Heat is added to change a solid to a liquid or a liquid to a gas; heat is removed to reverse these changes.

The **molar specific heat** of a gas depends on the process.

$$Q = \begin{cases} \pm ML_f & (\text{melt/freeze}) \\ \pm ML_v & (\text{boil/condense}) \end{cases}$$

$$\left\{ \begin{array}{l} \text{For a constant-volume process: } Q = nC_V \Delta T \\ \text{For a constant-pressure process: } Q = nC_P \Delta T \end{array} \right.$$

Mechanisms of heat transfer

An object can transfer heat to other objects or to its environment:

Conduction is the transfer of heat by direct physical contact.

$$\frac{Q}{\Delta t} = \left(\frac{kA}{L} \right) \Delta T$$

Convection is the transfer of heat by the motion of a fluid.

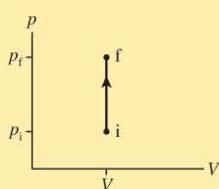


Radiation is the transfer of heat by electromagnetic waves.

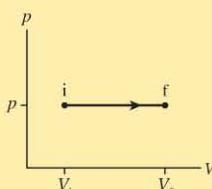
$$\frac{Q}{\Delta t} = e\sigma AT^4$$

A **pV (pressure-volume) diagram** is a useful means of looking at a process involving a gas.

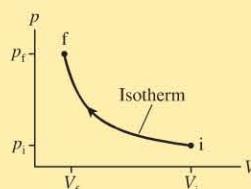
A **constant-volume** process has no change in volume.



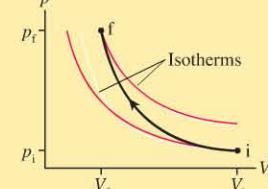
An **isobaric** process happens at a constant pressure.



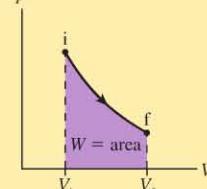
An **isothermal** process happens at a constant temperature.



An **adiabatic** process involves no transfer of heat; the temperature changes.



The work done by a gas is the area under the graph.



APPLICATIONS

Thermal expansion Objects experience an increase in volume and an increase in length when their temperature changes:

$$\Delta V = \beta V_i \Delta T \quad \Delta L = \alpha L_i \Delta T$$

Calorimetry When two or more systems interact thermally, they come to a common final temperature determined by

$$Q_{\text{net}} = Q_1 + Q_2 + Q_3 + \dots = 0$$

The number of **moles** is

$$n = \frac{M \text{ (in grams)}}{M_{\text{mol}}}$$



For homework assigned on MasteringPhysics, go to
www.masteringphysics.com

Problems labeled INT integrate significant material from earlier chapters; BIO are of biological or medical interest.

Problem difficulty is labeled as I (straightforward) to III (challenging).

VIEW ALL SOLUTIONS

QUESTIONS

Conceptual Questions

1. Which has more mass, a mole of Ne gas or a mole of N₂ gas?
 2. If you launch a projectile upward with a high enough speed, its kinetic energy is sufficient to allow it to escape the earth's gravity—it will go up and not come back down. Given enough time, hydrogen and helium gas atoms in the earth's atmosphere will escape, so these elements are not present in our atmosphere. Explain why hydrogen and helium atoms have the necessary speed to escape but why other elements, such as oxygen and nitrogen, do not.
 3. You may have noticed that latex helium balloons tend to shrink rather quickly; a balloon filled with air lasts a lot longer. Balloons shrink because gas diffuses out of them. The rate of diffusion is faster for smaller particles and for particles of higher speed. Diffusion is also faster when there is a large difference in concentration between two sides of a membrane. Given these facts, explain why an air-filled balloon lasts longer than a helium balloon.
 4. If you buy a sealed bag of potato chips in Miami and drive with it to Denver, where the atmospheric pressure is lower, you will find that the bag gets very "puffy." Explain why.
 5. If you double the typical speed of the molecules in a gas, by what factor does the pressure change? Give a simple explanation why the pressure changes by this factor.
 6. Two gases have the same number of molecules per cubic meter (N/V) and the same rms speed. The molecules of gas 2 are more massive than the molecules of gas 1.
 - a. Do the two gases have the same pressure? If not, which is larger?
 - b. Do the two gases have the same temperature? If not, which is larger?
 7. a. Which contains more particles, a mole of helium gas or a mole of oxygen gas? Explain.
 b. Which contains more particles, a gram of helium gas or a gram of oxygen gas? Explain.
 8. You have 100 g of aluminum and 100 g of lead.
 - a. Which contains a larger number of moles? Explain.
 - b. Which contains more atoms? Explain.
 9. Suppose you could suddenly increase the speed of every atom in a gas by a factor of 2.
 - a. Does the thermal energy of the gas change? If so, by what factor? If not, why not?
 - b. Does the molar specific heat change? If so, by what factor? If not, why not?
 10. A gas cylinder contains 1.0 mol of helium at a temperature of 20°C. A second identical cylinder contains 1.0 mol of neon at 20°C. The helium atoms are moving with a higher average speed, but the gas pressure in the two containers is the same. Explain how this is possible.
 11. A gas is in a sealed container. By what factor does the gas pressure change if
- a. The volume is doubled and the temperature is tripled?
 - b. The volume is halved and the temperature is tripled?
 12. A gas is in a sealed container. By what factor does the gas temperature change if
 - a. The volume is doubled and the pressure is tripled?
 - b. The volume is halved and the pressure is tripled?
 13. What is the maximum amount of work that a gas can do during a constant-volume process?
 14. You need to precisely measure the dimensions of a large wood panel for a construction project. Your metal tape measure was left outside for hours in the sun on a hot summer day, and now the tape is so hot it's painful to pick up. How will your measurements differ from those taken by your coworker, whose tape stayed in the shade? Explain.
 15. Your car's radiator is made of steel and is filled with water. You fill the radiator to the very top with cold water, then drive off without remembering to replace the cap. As the water and the steel radiator heat up, will the level of water drop or will it rise and overflow? Explain.
 16. Materials A and B have equal densities, but A has a larger specific heat than B. You have 100 g cubes of each material. Cube A, initially at 0°C, is placed in good thermal contact with cube B, initially at 200°C. The cubes are inside a well-insulated container where they don't interact with their surroundings. Is their final temperature greater than, less than, or equal to 100°C? Explain.
 17. Two containers hold equal masses of nitrogen gas at equal temperatures. You supply 10 J of heat to container A while not allowing its volume to change, and you supply 10 J of heat to container B while not allowing its pressure to change. Afterward, is temperature T_A greater than, less than, or equal to T_B ? Explain.
 18. You need to raise the temperature of a gas by 10°C. To use the smallest amount of heat energy, should you heat the gas at constant pressure or at constant volume? Explain.
 19. A sample of ideal gas is in a cylinder with a movable piston. 600 J of heat is added to the gas in an isothermal process. As the gas expands, pushing against the piston, how much work does it do?
 20. A student is heating chocolate in a pan on the stove. He uses a cooking thermometer to measure the temperature of the chocolate and sees it varies as shown in Figure Q12.20. Describe what is happening to the chocolate in each of the three portions of the graph.

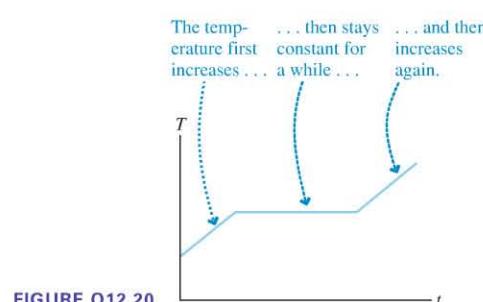


FIGURE Q12.20

21. If you bake a cake at high elevation, where atmospheric pressure is lower than at sea level, you will need to adjust the recipe. You will need to cook the cake for a longer time, and you will need to add less baking powder. (Baking powder is a leavening agent. As it heats, it releases gas bubbles that cause the cake to rise.) Explain why those adjustments are necessary.

22. The specific heat of aluminum is higher than that of iron. 1 kg blocks of iron and aluminum are heated to 100°C, and each is then dropped into its own 1 L beaker of 20°C water. Which beaker will end up with the warmer water? Explain.

23. A student is asked to sketch a pV diagram for a gas that goes through a cycle consisting of
 (a) an isobaric expansion, (b) a constant-volume reduction in temperature, and (c) an isothermal process that returns the gas to its initial state. The student draws the diagram shown in Figure Q12.23. What, if anything, is wrong with the student's diagram?

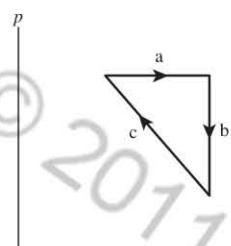


FIGURE Q12.23

24. If you have two spoons of the same size, one silver and one stainless steel, there is a quick test to tell which is which. Hold the end of a spoon in each hand, then lower them both into a cup of very hot water. One spoon will feel hot first. Is that the silver spoon or the stainless steel spoon? Explain.

25. If you live somewhere with cold, clear nights, you may have noticed some mornings when there was frost on open patches of ground but not under trees. This is because the ground under trees does not get as cold as open ground. Explain how tree cover keeps the ground under trees warmer.





FIGURE Q12.23

Multiple-Choice Questions

26. I A tire is inflated to a gauge pressure of 35 psi. The absolute pressure in the tire is

 - Less than 35 psi.
 - Equal to 35 psi.
 - Greater than 35 psi.

27. I The number of atoms in a container is increased by a factor of 2 while the temperature is held constant. The pressure

 - Decreases by a factor of 4.
 - Decreases by a factor of 2.
 - Stays the same.
 - Increases by a factor of 2.
 - Increases by a factor of 4.

28. II A gas is compressed by an isothermal process that decreases its volume by a factor of 2. In this process, the pressure

 - Does not change.
 - Increases by a factor of less than 2.
 - Increases by a factor of 2.
 - Increases by a factor of more than 2.

29. **III** A gas is compressed by an adiabatic process that decreases its volume by a factor of 2. In this process, the pressure

 - Does not change.
 - Increases by a factor of less than 2.
 - Increases by a factor of 2.
 - Increases by a factor of more than 2.

30. **I** Suppose you do a calorimetry experiment to measure the specific heat of a penny. You take a number of pennies, measure their mass, heat them to a known temperature, and then drop them into a container of water at a known temperature. You then deduce the specific heat of a penny by measuring the temperature change of the water. Unfortunately, you didn't realize that you dropped one penny on the floor while transferring them to the water. This will

 - Cause you to underestimate the specific heat.
 - Cause you to overestimate the specific heat.
 - Not affect your calculation of specific heat.

31. **I** A cup of water is heated with a heating coil that delivers 100 W of heat. In one minute, the temperature of the water rises by 20°C . What is the mass of the water?

 - 72 g
 - 140 g
 - 720 g
 - 1.4 kg

32. **I** Three identical beakers each hold 1000 g of water at 20°C . 100 g of liquid water at 0°C is added to the first beaker, 100 g of ice at 0°C is added to the second beaker, and the third beaker gets 100 g of aluminum at 0°C . The contents of which container end up at the lowest final temperature?

 - The first beaker.
 - The second beaker.
 - The third beaker.
 - All end up at the same temperature.

33. **III** 100 g of ice at 0°C and 100 g of steam at 100°C interact thermally in a well-insulated container. The final state of the system is

 - An ice-water mixture at 0°C .
 - Water at a temperature between 0°C and 50°C .
 - Water at 50°C .
 - Water at a temperature between 50°C and 100°C .
 - A water-steam mixture at 100°C .

34. **I** Suppose the 600 W of radiation emitted in a microwave oven is absorbed by 250 g of water in a very lightweight cup. Approximately how long will it take to heat the water from 20°C to 80°C ?

 - 50 s
 - 100 s
 - 150 s
 - 200 s

35. **II** 40,000 J of heat is added to 1.0 kg of ice at -10°C . How much ice melts?

 - 0.012 kg
 - 0.057 kg
 - 0.12 kg
 - 1.0 kg

36. **II** Steam at 100°C causes worse burns than liquid water at **BIO** 100°C . This is because

 - The steam is hotter than the water.
 - Heat is transferred to the skin as steam condenses.
 - Steam has a higher specific heat than water.

VIEW ALL SOLUTIONS

PROBLEMS

Section 12.1 The Atomic Model of Matter

1. | Which contains the most moles: 10 g of hydrogen gas, 100 g of carbon, or 50 g of lead?
2. ||| How many grams of water (H_2O) have the same number of oxygen atoms as 1.0 mol of oxygen gas?
3. ||| How many atoms of hydrogen are in 100 g of hydrogen peroxide (H_2O_2)?
4. || How many cubic millimeters (mm^3) are in 1 L?
5. || A box is 200 cm wide, 40 cm deep, and 3.0 cm high. What is its volume in m^3 ?

Section 12.2 The Atomic Model of an Ideal Gas

6. ||| Dry ice is frozen carbon dioxide. If you have 1.0 kg of dry ice, what volume will it occupy if you heat it enough to turn it into a gas at a temperature of 20°C?
7. | What is the absolute pressure of the air in your car's tires, in psi, when your pressure gauge indicates they are inflated to 35.0 psi? Assume you are at sea level.
8. ||| Total lung capacity of a typical adult is approximately 5.0 L.
- BIO Approximately 20% of the air is oxygen. At sea level and at an average body temperature of 37°C, how many moles of oxygen do the lungs contain at the end of an inhalation?
9. ||| Many cultures around the world still use a simple weapon called a blowgun, a tube with a dart that fits tightly inside. A sharp breath into the end of the tube launches the dart. When exhaling forcefully, a healthy person can supply air at a gauge pressure of 6.0 kPa. What force does this pressure exert on a dart in a 1.5-mm-diameter tube?
10. ||| When you stifle a sneeze, you can damage delicate tissues because the pressure of the air that is not allowed to escape may rise by up to 45 kPa. If this extra pressure acts on the inside of your 8.4-mm-diameter eardrum, what is the outward force?
11. ||| 7.5 mol of helium are in a 15 L cylinder. The pressure gauge on the cylinder reads 65 psi. What are (a) the temperature of the gas in °C and (b) the average kinetic energy of a helium atom?
12. || Mars has an atmosphere composed almost entirely of carbon dioxide, with an average temperature of -63°C. What is the rms speed of a molecule in Mars's atmosphere?
13. || 3.0 mol of gas at a temperature of -120°C fills a 2.0 L container. What is the gas pressure?
14. | 265 m/s is a typical cruising speed for a jet airliner. At what temperature (in °C) do the molecules of nitrogen gas have an rms speed of 265 m/s?
15. ||| 10 g of liquid water is placed in a flexible bag, the air is excluded, and the bag is sealed. It is then placed in a microwave oven where the water is boiled to make steam at 100°C. What is the volume of the bag after all the water has boiled? Assume that the pressure inside the bag is equal to atmospheric pressure.

Section 12.3 Ideal-Gas Processes

16. || A cylinder contains 3.0 L of oxygen at 300 K and 2.4 atm. The gas is heated, causing a piston in the cylinder to move outward. The heating causes the temperature to rise to 600 K and the volume of the cylinder to increase to 9.0 L. What is the final gas pressure?
17. ||| A gas with initial conditions p_i , V_i , and T_i expands isothermally until $V_f = 2V_i$. What are (a) T_f and (b) p_f ?

18. ||| 0.10 mol of argon gas is admitted to an evacuated 50 cm^3 container at 20°C. The gas then undergoes heating at constant volume to a temperature of 300°C.

- What is the final pressure of the gas?
- Show the process on a pV diagram. Include a proper scale on both axes.

19. | 0.10 mol of argon gas is admitted to an evacuated 50 cm^3 container at 20°C. The gas then undergoes an isobaric heating to a temperature of 300°C.

- What is the final volume of the gas?
- Show the process on a pV diagram. Include a proper scale on both axes.

20. ||| 0.10 mol of argon gas is admitted to an evacuated 50 cm^3 container at 20°C. The gas then undergoes an isothermal expansion to a volume of 200 cm^3 .

- What is the final pressure of the gas?
- Show the process on a pV diagram. Include a proper scale on both axes.

21. ||| 0.0040 mol of gas undergoes the process shown in Figure P12.21.

- What type of process is this?
- What are the initial and final temperatures?

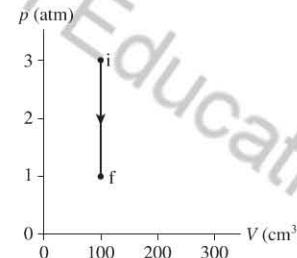


FIGURE P12.21

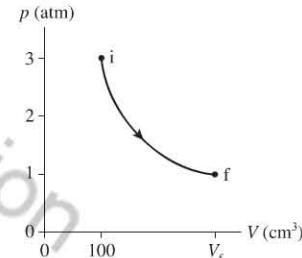


FIGURE P12.22

22. ||| 0.0040 mol of gas follows the hyperbolic trajectory shown in Figure P12.22.

- What type of process is this?
- What are the initial and final temperatures?
- What is the final volume V_f ?

23. || A gas with an initial temperature of 900°C undergoes the process shown in Figure P12.23.

- What type of process is this?
- What is the final temperature?
- How many moles of gas are there?

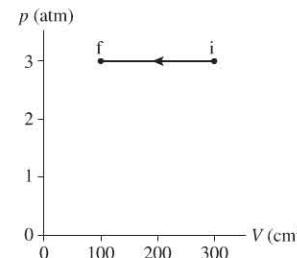
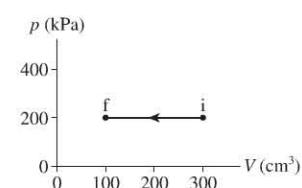


FIGURE P12.23



24. ||| How much work is done on the gas in the process shown in Figure P12.24?

25. || It is possible to make a thermometer by sealing gas in a rigid container and measuring the absolute pressure. Such a constant-volume gas thermometer is placed in an ice-water bath at 0.00°C . After reaching thermal equilibrium, the gas pressure is recorded as 55.9 kPa. The thermometer is then placed in contact with a sample of unknown temperature. After the thermometer reaches a new equilibrium, the gas pressure is 65.1 kPa. What is the temperature of this sample?
26. || A 1.0 cm^3 air bubble is released from the sandy bottom of a warm, shallow sea, where the gauge pressure is 1.5 atm. The bubble rises slowly enough that the air inside remains at the same constant temperature as the water.
- What is the volume of the bubble as it reaches the surface?
 - As the bubble rises, is heat energy transferred from the water to the bubble or from the bubble to the water? Explain.
27. || A weather balloon rises through the atmosphere, its volume expanding from 4.0 m^3 to 12 m^3 as the temperature drops from 20°C to -10°C . If the initial gas pressure inside the balloon is 1.0 atm, what is the final pressure?

Section 12.4 Thermal Expansion

28. || A straight rod consists of a 1.2-cm-long piece of aluminum attached to a 2.0-cm-long piece of steel. By how much will the length of this rod change if its temperature is increased from 20°C to 40°C ?
29. || The length of a steel beam increases by 0.73 mm when its temperature is raised from 22°C to 35°C . What is the length of the beam at 22°C ?
30. || Older railroad tracks in the U.S. are made of 12-m-long pieces of steel. When the tracks are laid, gaps are left between the sections to prevent buckling when the steel thermally expands. If a track is laid at 16°C , how large should the gaps be if the track is not to buckle when the temperature is as high as 50°C ?
31. |||| The temperature of an aluminum disk is increased by 120°C . By what percentage does its volume increase?

Section 12.5 Specific Heat and Heat of Transformation

32. ||| How much energy must be removed from a 200 g block of ice to cool it from 0°C to -30°C ?
33. ||| How much heat is needed to change 20 g of mercury at 20°C into mercury vapor at the boiling point?
34. || a. 100 J of heat energy are transferred to 20 g of mercury. By how much does the temperature increase?
b. How much heat is needed to raise the temperature of 20 g of water by the same amount?
35. || The maximum amount of water an adult in temperate climates can perspire in one hour is typically 1.8 L. However, after several weeks in a tropical climate the body can adapt, increasing the maximum perspiration rate to 3.5 L/h. At what rate, in watts, is energy being removed when perspiring that rapidly? Assume all of the perspired water evaporates. At body temperature, the heat of vaporization of water is $L_v = 24 \times 10^5 \text{ J/kg}$.
36. || Alligators and other reptiles don't use enough metabolic energy to keep their body temperatures constant. They cool off at night and must warm up in the sun in the morning. Suppose a

300 kg alligator with an early-morning body temperature of 25°C is absorbing radiation from the sun at a rate of 1200 W. How long will the alligator need to warm up to a more favorable 30°C ? (Assume that the specific heat of the reptilian body is the same as that of the mammalian body.)



37. || When air is inhaled, it quickly becomes saturated with water **BIO** vapor as it passes through the moist upper airways. When breathing dry air, about 25 mg of water are exhaled with each breath. At 12 breaths/min, what is the rate of energy loss due to evaporation? Express your answer in both watts and Calories per day. At body temperature, the heat of vaporization of water is $L_v = 24 \times 10^5 \text{ J/kg}$.
38. ||| It is important for the body to have mechanisms to effectively **BIO** cool itself; if not, moderate exercise could easily increase body temperatures to dangerous levels. Suppose a 70 kg man runs on a treadmill for 30 min, using a metabolic power of 1000 W. Assume that all of this power goes to thermal energy in the body. If he couldn't perspire or otherwise cool his body, by how much would his body temperature rise during this exercise?
39. ||| What minimum heat is needed to bring 100 g of water at 20°C to the boiling point and completely boil it away?

Section 12.6 Calorimetry

40. ||| 30 g of copper pellets are removed from a 300°C oven and immediately dropped into 100 mL of water at 20°C in an insulated cup. What will the new water temperature be?
41. ||| A copper block is removed from a 300°C oven and dropped into 1.00 kg of water at 20.0°C . The water quickly reaches 25.5°C and then remains at that temperature. What is the mass of the copper block?
42. ||| A 750 g aluminum pan is removed from the stove and plunged into a sink filled with 10.0 kg of water at 20.0°C . The water temperature quickly rises to 24.0°C . What was the initial temperature of the pan?
43. ||| A 500 g metal sphere is heated to 300°C , then dropped into a beaker containing 4.08 kg of mercury at 20.0°C . A short time later the mercury temperature stabilizes at 99.0°C . Identify the metal.
44. ||| Brewed coffee is often too hot to drink right away. You can cool it with an ice cube, but this dilutes it. Or you can buy a device that will cool your coffee without dilution—a 200 g aluminum cylinder that you take from your freezer and place in a mug of hot coffee. If the cylinder is cooled to -20°C , a typical freezer temperature, and then dropped into a large cup of coffee (essentially water, with a mass of 500 g) at 85°C , what is the final temperature of the coffee?
45. |||| Marianne really likes coffee, but on summer days she doesn't want to drink a hot beverage. If she is served 200 mL of coffee at 80°C in a well-insulated container, how much ice at 0°C should she add to obtain a final temperature of 30°C ?
46. || If a person has a dangerously high fever, submerging her in **BIO** ice water is a bad idea, but an ice pack can help to quickly bring her body temperature down. How many grams of ice at 0°C will be melted in bringing down a 60 kg patient's fever from 40°C to 39°C ?

Section 12.7 Thermal Properties of Gases

47. I A container holds 1.0 g of argon at a pressure of 8.0 atm.
- How much heat is required to increase the temperature by 100°C at constant volume?
 - How much will the temperature increase if this amount of heat energy is transferred to the gas at constant pressure?
48. II A container holds 1.0 g of oxygen at a pressure of 8.0 atm.
- How much heat is required to increase the temperature by 100°C at constant pressure?
 - How much will the temperature increase if this amount of heat energy is transferred to the gas at constant volume?
49. I What is the temperature change of 1.0 mol of a monatomic gas if its thermal energy is increased by 1.0 J?
50. II The temperature of 2.0 g of helium is increased at constant volume by ΔT . What mass of oxygen can have its temperature increased by the same amount at constant volume using the same amount of heat?
51. III How much work is done per cycle by a gas following the pV trajectory of Figure P12.51?

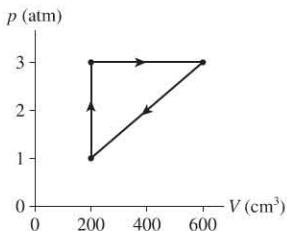


FIGURE P12.51

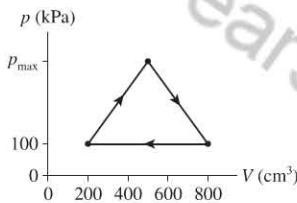
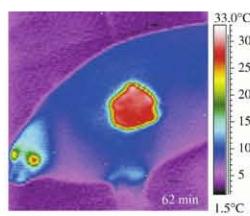


FIGURE P12.52

52. III A gas following the pV trajectory of Figure P12.52 does 60 J of work per cycle. What is p_{\max} ?

Section 12.8 Heat Transfer

53. III A 1.8-cm-thick wood floor covers a $4.0 \text{ m} \times 5.5 \text{ m}$ room. The subfloor on which the flooring sits is at a temperature of 16.2°C, while the air in the room is at 19.6°C. What is the rate of heat conduction through the floor?
54. IIII A copper-bottomed kettle, its bottom 24 cm in diameter and 3.0 mm thick, sits on a burner. The kettle holds boiling water, and energy flows into the water from the kettle bottom at 800 W. What is the temperature of the bottom surface of the kettle?
55. IIII What is the greatest possible rate of energy transfer by radiation from a metal cube 2.0 cm on a side that is at 700°C? Its emissivity is 0.20.
56. III What is the greatest possible rate of energy transfer by radiation for a 5.0-cm-diameter sphere that is at 100°C?
57. III Seals may cool themselves by using *thermal windows*, patches on their bodies with much higher than average surface temperature. Suppose a seal has a 0.030 m^2 thermal window at a temperature of 30°C. If the seal's surroundings are a frosty -10°C, what is the net rate of energy loss by radiation? Assume an emissivity equal to that of a human.
58. II Electronics and inhabitants of the International Space Station generate a significant amount of thermal energy that the station must get rid of. The only way that the station can exhaust thermal energy is by radiation, which it does using thin, 1.8-m-by-3.6-m



panels that have a working temperature of about 6°C. How much power is radiated from each panel? Assume that the panels are in the shade so that the absorbed radiation will be negligible. Assume that the emissivity of the panels is 1.0.

Hint: Don't forget that the panels have two sides!

59. II The glowing filament in a lamp is radiating energy at a rate of 60 W. At the filament's temperature of 1500°C, the emissivity is 0.23. What is the surface area of the filament?
60. III If you lie on the ground at night with no cover, you get cold rather quickly. Much of this is due to energy loss by radiation. At night in a dry climate, the temperature of the sky can drop to -40°C. If you are lying on the ground with thin clothing that provides little insulation, the surface temperature of your skin and clothes will be about 30°C. Estimate the net rate at which your body loses energy by radiation to the night sky under these conditions. **Hint:** What area should you use?

General Problems

61. II A rigid container holds 2.0 mol of gas at a pressure of 1.0 atm and a temperature of 30°C.
- What is the container's volume?
 - What is the pressure if the temperature is raised to 130°C?
62. II A 15-cm-diameter compressed-air tank is 50 cm tall. The pressure at 20°C is 150 atm.
- How many moles of air are in the tank?
 - What volume would this air occupy at STP?
63. IIII A 10-cm-diameter cylinder of helium gas is 30 cm long and at 20°C. The pressure gauge reads 120 psi.
- How many helium atoms are in the cylinder?
 - What is the mass of the helium?
64. II Party stores sell small tanks containing 30 g of helium gas. If you use such a tank to fill 0.010 m^3 foil balloons (which don't stretch, and so have an internal pressure that is very close to atmospheric pressure), how many balloons can you expect to fill?
65. IIII Suppose you take and hold a deep breath on a chilly day, **BIO** inhaling 3.0 L of air at 0°C. Assume that air pressure is 1.0 atm.
- How much heat must your body supply to warm the air to your internal body temperature of 37°C?
 - How much does the air's volume increase as it is warmed?
66. II On average, each person in the industrialized world is responsible for the emission of 10,000 kg of carbon dioxide (CO_2) every year. This includes CO_2 that you generate directly, by burning fossil fuels to operate your car or your furnace, as well as CO_2 generated on your behalf by electric generating stations and manufacturing plants. CO_2 is a greenhouse gas that contributes to global warming. If you were to store your yearly CO_2 emissions in a cube at STP, how long would each edge of the cube be?
67. III On a cool morning, when the temperature is 15°C, you measure the pressure in your car tires to be 30 psi. After driving 20 mi on the freeway, the temperature of your tires is 45°C. What pressure will your tire gauge now show?
68. III Suppose you inflate your car tires to 35 psi on a 20°C day. Later, the temperature drops to 0°C. What is the pressure in your tires now?
69. I The volume in a constant-pressure gas thermometer is directly proportional to the absolute temperature. A constant-pressure thermometer is calibrated by adjusting its volume to 1000 mL while it is in contact with a reference cell at 0.01°C. The volume increases to 1638 mL when the thermometer is placed in contact with a sample. What is the sample's temperature in °C?

70. II A compressed-air cylinder is known to fail if the pressure exceeds 110 atm. A cylinder that was filled to 25 atm at 20°C is stored in a warehouse. Unfortunately, the warehouse catches fire and the temperature reaches 950°C. Does the cylinder explode?
71. II A rigid sphere has a valve that can be opened or closed. The sphere with the valve open is placed in boiling water in a room where the air pressure is 1.0 atm. After a long period of time has elapsed, the valve is closed. What will be the pressure inside the sphere if it is then placed in (a) a mixture of ice and water and (b) an insulated box filled with dry ice, which is at -78.5°C?
72. III 80 J of work are done on the gas in the process shown in Figure P12.72. What is V_f in cm³?

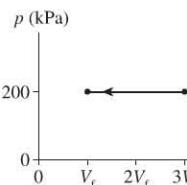


FIGURE P12.72

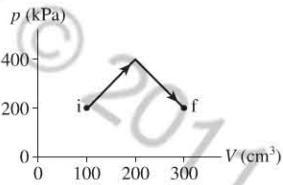


FIGURE P12.73

73. III How much work is done by the gas in the process shown in Figure P12.73?
74. I 0.10 mol of gas undergoes the process 1 → 2 shown in Figure P12.74.
- What are temperatures T_1 and T_2 ?
 - What type of process is this?
 - The gas undergoes constant-volume heating from point 2 until the pressure is restored to the value it had at point 1. What is the final temperature of the gas?

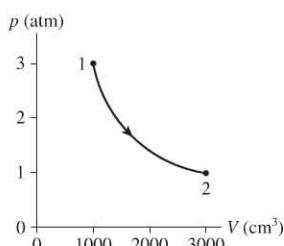


FIGURE P12.74

75. II 10 g of dry ice (solid CO₂) is placed in a 10,000 cm³ container, then all the air is quickly pumped out and the container sealed. The container is warmed to 0°C, a temperature at which CO₂ is a gas.
- What is the gas pressure? Give your answer in atm.
 - The gas then undergoes an isothermal compression until the pressure is 3.0 atm, immediately followed by an isobaric compression until the volume is 1000 cm³.
 - What is the final temperature of the gas?
 - Show the process on a pV diagram.

76. III A large freshwater fish has a swim bladder with a volume of BIO 5.0 × 10⁻⁴ m³. The fish descends from a depth where the absolute pressure is 3.0 atm to deeper water where the swim bladder is compressed to 60% of its initial volume. As the fish descends, the gas pressure in the swim bladder is always equal to the water pressure, and the temperature of the gas remains at the internal temperature of the fish's body. To adapt to its new location, the fish must add gas to reinflate its swim bladder to the original volume. This takes energy to accomplish. What's the minimum amount of work required to expand the swim bladder back to its original volume?

77. III A 5.0-m-diameter garden pond holds 5.9 × 10³ kg of water. Solar energy is incident on the pond at an average rate of 400 W/m². If the water absorbs all the solar energy and does not exchange energy with its surroundings, how many hours will it take to warm from 15°C to 25°C?

78. III 0.030 mol of an ideal monatomic gas undergoes an adiabatic compression that raises its temperature from 10°C to 50°C. How much work is done on the gas to compress it?
79. III 0.15 mol of an ideal monatomic gas undergoes an adiabatic expansion, cooling from 20°C to -10°C. How much work is done by the gas during the expansion?

80. III Susan, whose mass is 68 kg, climbs 59 m to the top of the BIO Cape Hatteras lighthouse.

- INT a. During the climb, by how much does her potential energy increase?
- b. For a typical efficiency of 25%, what metabolic energy does she require to complete the climb?
- c. When exercising, the body must perspire and use other mechanisms to cool itself to avoid potentially dangerous increases on body temperature. If we assume that Susan doesn't perspire or otherwise cool herself and that all of the "lost" energy goes into increasing her body temperature, by how much would her body temperature increase during this climb?

81. III A typical nuclear reactor generates 1000 MW of electric energy. In doing so, it produces "waste heat" at a rate of 2000 MW, and this heat must be removed from the reactor. Many reactors are sited next to large bodies of water so that they can use the water for cooling. Consider a reactor where the intake water is at 18°C. State regulations limit the temperature of the output water to 30°C so as not to harm aquatic organisms. How many kilograms of cooling water have to be pumped through the reactor each minute?

82. III A 68 kg woman cycles at a constant 15 km/h. All of the metabolic energy that does not go to forward propulsion is converted INT to thermal energy in her body. If the only way her body has to keep cool is by evaporation, how many kilograms of water must she lose to perspiration each hour to keep her body temperature constant?

83. III A 1200 kg car traveling at 60 mph quickly brakes to a halt. INT The kinetic energy of the car is converted to thermal energy of the disk brakes. The brake disks (one per wheel) are iron disks with a mass of 4.0 kg. Estimate the temperature rise in each disk as the car stops.

84. III A 5000 kg African elephant BIO has a resting metabolic rate of INT 2500 W. On a hot day, the elephant's environment is likely to be nearly the same temperature as the animal itself, so cooling by radiation is not effective. The only plausible way to keep cool is by evaporation, and elephants spray water on their body to accomplish this. If this is the only possible means of cooling, how many kilograms of water per hour must be evaporated from an elephant's skin to keep it at a constant temperature?



85. III Suppose you drop a water balloon from a height of 10 m. If the balloon doesn't break on impact, its kinetic energy will be converted to thermal energy. Estimate the temperature rise of the water. Is this likely to be noticeable?

86. || What is the maximum mass of lead you could melt with 1000 J of heat, starting from 20°C?
87. || An experiment measures the temperature of a 200 g substance while steadily supplying heat to it. Figure P12.87 shows the results of the experiment. What are (a) the specific heat of the liquid phase and (b) the heat of vaporization?

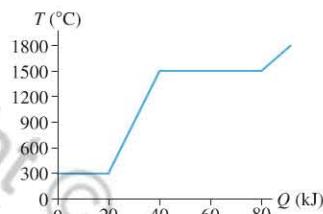


FIGURE P12.87

88. ||| 10 g of aluminum at 200°C and 20 g of copper are dropped into 50 cm³ of ethyl alcohol at 15°C. The temperature quickly comes to 25°C. What was the initial temperature of the copper?
89. ||| A 100 g ice cube at -10°C is placed in an aluminum cup whose initial temperature is 70°C. The system comes to an equilibrium temperature of 20°C. What is the mass of the cup?
90. ||| A 50.0 g thermometer is used to measure the temperature of 200 g of water. The specific heat of the thermometer, which is mostly glass, is 750 J/kg · K, and it reads 20.0°C while lying on the table. After being completely immersed in the water, the thermometer's reading stabilizes at 71.2°C. What was the actual water temperature before it was measured?
91. ||| Your 300 mL cup of coffee is too hot to drink when served at 90°C. What is the mass of an ice cube, taken from a -20°C freezer, that will cool your coffee to a pleasant 60°C?
92. || A gas is compressed from 600 cm³ to 200 cm³ at a constant pressure of 400 kPa. At the same time, 100 J of heat energy is transferred out of the gas. What is the change in thermal energy of the gas during this process?
93. ||| An expandable cube, initially 20 cm on each side, contains 3.0 g of helium at 20°C. 1000 J of heat energy are transferred to this gas. What are (a) the final pressure if the process is at constant volume and (b) the final volume if the process is at constant pressure?
94. || 0.10 mol of a monatomic gas follows the process shown in Figure P12.94.
- How much heat energy is transferred to or from the gas during process 1 → 2?
 - How much heat energy is transferred to or from the gas during process 2 → 3?
 - What is the total change in thermal energy of the gas?

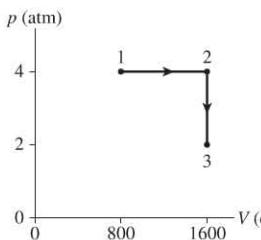


FIGURE P12.94

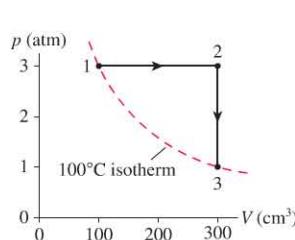


FIGURE P12.95

95. || A monatomic gas follows the process 1 → 2 → 3 shown in Figure P12.95. How much heat is needed for (a) process 1 → 2 and (b) process 2 → 3?

96. ||| What are (a) the heat Q_H extracted from the hot reservoir and (b) the efficiency for a heat engine described by the pV diagram of Figure P12.96?

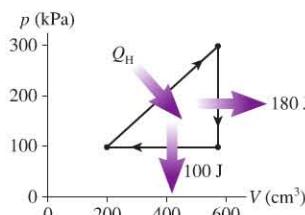


FIGURE P12.96

97. ||| The top layer of your goose down sleeping bag has a thickness of 5.0 cm and a surface area of 1.0 m². When the outside temperature is -20°C, you lose 25 Cal/h by heat conduction through the bag (which remains at a cozy 35°C inside). Assume that you're sleeping on an insulated pad that eliminates heat conduction to the ground beneath you. What is the thermal conductivity of the goose down?
98. ||| Suppose you go outside in your fiber-filled jacket on a windless but very cold day. The thickness of the jacket is 2.5 cm, and it covers 1.1 m² of your body. The purpose of fiber- or down-filled jackets is to trap a layer of air, and it's really the air layer that provides the insulation. If your skin temperature is 34°C while the air temperature is -20°C, at what rate is heat being conducted through the jacket and away from your body?
99. ||| Two thin, square copper plates are radiating energy at the same rate. The edge length of plate 2 is four times that of plate 1. What is the ratio of absolute temperatures T_1 / T_2 of the plates?
100. ||| The surface area of an adult human is about 1.8 m². Suppose a person with a skin temperature of 34°C is standing with bare skin in a room where the air is 25°C but the walls are 17°C.
- There is a "dead-air" layer next to your skin that acts as insulation. If the dead-air layer is 5.0 mm thick, what is the person's rate of heat loss by conduction?
 - What is the person's net radiation loss to the walls? The emissivity of skin is 0.97.
 - Does conduction or radiation contribute more to the total rate of energy loss?
 - If the person is metabolizing food at a rate of 155 W, does the person feel comfortable, chilly, or too warm?

Passage Problems

Thermal Properties of the Oceans

Seasonal temperature changes in the ocean only affect the top layer of water, to a depth of 500 m or so. This "mixed" layer is thermally isolated from the cold, deep water below. The average temperature of this top layer of the world's oceans, which has area 3.6×10^8 km², is approximately 17°C.

In addition to seasonal temperature changes, the oceans have experienced an overall warming trend over the last century that is expected to continue as the earth's climate changes. A warmer ocean means a larger volume of water; the oceans will rise. Suppose the average temperature of the top layer of the world's oceans were to increase from a temperature T_i to a temperature T_f . The area of the oceans will not change, as this is fixed by the size of the ocean basin,

so any thermal expansion of the water will cause the water level to rise, as shown in Figure P12.101. The original volume is the product of the original depth and the surface area, $V_i = Ad_i$. The change in volume is given by $\Delta V = A \Delta d$.

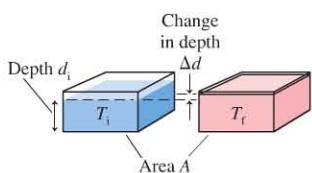


FIGURE P12.101

101. | If the top 500 m of ocean water increased in temperature from 17°C to 18°C, what would be the resulting rise in ocean height?
 A. 0.11 m B. 0.22 m
 C. 0.44 m D. 0.88 m
102. | Approximately how much energy would be required to raise the temperature of the top layer of the oceans by 1°C? (1 m³ of water has a mass of 1000 kg.)
 A. 1×10^{24} J B. 1×10^{21} J
 C. 1×10^{18} J D. 1×10^{15} J

103. | Water's coefficient of expansion varies with temperature. For water at 2°C, an increase in temperature of 1°C would cause the volume to

A. Increase. B. Stay the same. C. Decrease.

104. | The ocean is mostly heated from the top, by light from the sun. The warmer surface water doesn't mix much with the colder deep ocean water. This lack of mixing can be ascribed to a lack of
 A. Conduction.
 B. Convection.
 C. Radiation.
 D. Evaporation.

STOP TO THINK ANSWERS

Stop to Think 12.1: A. Both helium and neon are monatomic gases, where the basic particles are atoms. 5 mol of helium contain 5 times as many atoms as 1 mol of neon, though both samples have the same mass.

Stop to Think 12.2: A, B. An increase in temperature means that the atoms have a larger average kinetic energy and will thus have a larger rms speed. Because the thermal energy of the gas is simply the total kinetic energy of the atoms, this must increase as well. The pressure *could* change, but it's not required; T and V could increase by the same factor, which would keep p constant. The container is sealed, so the number of molecules does not change.

Stop to Think 12.3: B. The product pV/T is constant. During the process, pV decreases by a factor of 2, so T must decrease by a factor of 2 as well.

Stop to Think 12.4: A. The thermal expansion coefficients of aluminum are greater than those of iron. Heating the rod and the ring will expand the outer diameter of the rod and the inner diameter of the ring, but the ring's expansion will be greater.

Stop to Think 12.5: B. To solidify the lead, heat must be removed; this heat boils the water. The heat of vaporization of water is 10 times the heat of fusion of lead, so much less than 1 kg of water vaporizes as 1 kg of lead solidifies.

Stop to Think 12.6: A. The lead cools and the water warms as heat is transferred from the lead to the water. The specific heat of water is much larger than that of lead, so the temperature change of the water is much less than that of the lead.

Stop to Think 12.7: C. With a sealed suit and no matter around you, there is no way to transfer heat to the environment except by radiation.