form of sound. The region near the border between infrared and radio waves, where wavelengths range from micrometers to millimeters, is sometimes given the name **microwaves.**

On the other side of the spectrum, light with wavelengths somewhat shorter than blue light is called **ultraviolet**, because it lies beyond the blue (or violet) end of the rainbow. Light with even shorter wavelengths is called **X rays**, and the shortest-wavelength light is called **gamma rays.** Notice that visible light is an extremely small part of the entire electromagnetic spectrum: The reddest red that our eyes can see has only about twice the wavelength of the bluest blue, but the radio waves from your favorite radio station are a billion times longer than the X rays used in a doctor's office.

The different energies of different forms of light explain many familiar effects in everyday life. Radio waves carry so little energy that they have no noticeable effect on our bodies. However, radio waves can make electrons move up and down in an antenna, which is how your car radio receives the radio waves coming from a radio station. Molecules moving around in a warm object emit infrared light, which is why we sometimes associate infrared light with heat. Receptors in our eyes respond to visible-light photons, making vision possible. Ultraviolet photons carry enough energy to harm cells in our skin, causing sunburn or skin cancer. X-ray photons have enough energy to penetrate through skin and muscle but can be blocked by bones or teeth. That is why doctors and dentists can see our bone structures on photographs taken with X-ray light.

Just as different colors of visible light may be absorbed or reflected differently by the objects we see (see Figure 5.3), the various portions of the electromagnetic spectrum may interact in very different ways with matter. For example, a brick wall is opaque to visible light but transmits radio waves, which is why radios and cell phones work inside buildings. Similarly, glass that is transparent to visible light may be opaque to ultraviolet light. In general, certain types of matter tend to interact more strongly with certain types of light, so each type of light carries different information about distant objects in the universe. That is why astronomers seek to observe light of all wavelengths, from radio waves to gamma rays [Section 6.4].

COMMON MISCONCEPTIONS

What Could You See with X-Ray Vision?

The idea of having "X-ray vision" that would allow you to see through walls is part of Hollywood culture and dates back to old Superman comics. But is it really possible? At first, the idea might seem reasonable; after all, you've probably seen a doctor or dentist holding up "X rays" that allow you to see your bones or teeth. However, the "X rays" that the doctor or dentist holds up are not a form of light; they are just pieces of film. The images of your bones and teeth are made with the help of a special machine that works somewhat like the flash on an ordinary camera but emits X rays instead of visible light. This machine flashes the X rays at you, and the X rays that are transmitted through your body are recorded on film or with an electronic detector. Thus, what you see is the image left by these X rays, not the X rays themselves, which have wavelengths far too short for human eyes. If you think about this process, you'll realize that X-ray vision would do you no good, even if it were possible. People, walls, and other ordinary objects do not emit any X rays of their own, so there'd be nothing to see with your X-ray vision.

COMMON MISCONCEPTIONS

Is Radiation Dangerous?

Many people associate the word *radiation* with danger. However, the word *radiate* simply means "to spread out from a center" (note the similarity between *radiation* and *radius* [of a circle]). Radiation is energy being carried through space. If energy is being carried by particles of matter, such as protons or neutrons, we call it *particle radiation*. If energy is being carried by light, we call it *electromagnetic radiation*.

High-energy forms of radiation, such as particles from radioactive materials or X rays, are dangerous because they can penetrate body tissues and cause cell damage. Low-energy forms of radiation, such as radio waves, are usually harmless. The visible-light radiation from the Sun is necessary to life on Earth. Thus, while some forms of radiation are dangerous, others are harmless or beneficial.

5.3 PROPERTIES OF MATTER

Light carries information about matter across the universe, but we are usually more interested in the matter the light is coming from than in the light itself. Planets, stars, and galaxies are made of matter, and we must understand the nature of matter if we are to decode the messages we receive in light.

What is the structure of matter?

Like the nature of light, the nature of matter remained mysterious for most of human history. Nevertheless, ancient philosophers came up with some ideas that are still with us today.

The ancient Greek philosopher Democritus (c. 470–380 B.C.) wondered what would happen if we broke a piece of matter, such as a rock, into ever smaller pieces. Democritus believed that the rock would eventually break into particles so small that nothing smaller could be possible. He called these particles *atoms*, a Greek term meaning "indivisible." Building upon the beliefs of earlier Greek philosophers, Democritus thought that all materials were composed of just four basic *elements:* fire, water, earth, and air. He proposed that the different properties of the elements could be explained by the physical characteristics of their atoms. For example, Democritus suggested that atoms of water were smooth and round, so water flowed and had no fixed shape, while burns were painful because atoms of fire were thorny. He also imagined atoms of earth to be rough and jagged, so that they could fit together like pieces of a three-dimensional jigsaw puzzle, and he used this idea to suggest that the universe began as a chaotic mix of atoms that slowly clumped together to form Earth.

Although Democritus was wrong about there being only four types of atoms and about their specific properties, he was on the right track. All ordinary matter is indeed composed of atoms, and the properties of ordinary matter depend on the physical characteristics of their atoms. However, by modern definition, atoms are *not* indivisible because they are composed of even smaller particles.

Atoms come in different types, and each type corresponds to a different chemical **element**. Today, we have identified more than 100 chemical elements, and fire, water, earth, and air are *not* among them. Some of the most familiar chemical elements are hydrogen, helium, carbon, oxygen, silicon, iron, gold, silver, lead, and uranium. Appendix D gives the periodic table of all the elements.

Atomic Structure Each chemical element consists of a different type of atom, and atoms are in turn made of particles that we call **protons**, **neutrons**, and **electrons** (Figure 5.8). Protons and neutrons are found in the tiny **nucleus** at the

MATHEMATICAL INSIGHT 5.1

Wavelength, Frequency, and Energy

The speed of any wave is the product of its wavelength and its frequency. Because all forms of light travel at the same speed (in a vacuum) of $c = 3 \times 10^8$ m/s, we can write

$$\lambda \times f = c$$

where λ (the Greek letter *lambda*) stands for wavelength and f stands for frequency. This formula is simple but revealing: Because the speed c is constant, the formula tells us that frequency must go up when wavelength goes down, and vice versa.

The formula for the radiative energy (E) carried by a photon of light is

$$E = h \times f$$

where h is *Planck's constant* ($h = 6.626 \times 10^{-34}$ joule $\times$ s). Thus, energy increases in proportion to the frequency of the photon.

EXAMPLE 1: The numbers on a radio dial for FM radio stations are the frequencies of radio waves in megahertz (MHz), or millions of hertz. If your favorite radio station is "93.3 on your dial," it broadcasts radio waves with a frequency of 93.3 million cycles per second. What is the wavelength of these radio waves?

SOLUTION:

Step 1 Understand: Radio waves are a form of light, so we know they obey the relationship

$$\lambda \times f = c$$

The speed of light is always the same and we are given the frequency ($f = 93.3$ MHz), so we have all the information needed to solve for the wavelength (λ).

Step 2 Solve: We solve for wavelength by dividing both sides of the above equation by the frequency (f), finding

$$\lambda = \frac{c}{f}$$

Now we plug in the speed of light (3×10^8 m/s) and the frequency. Remember that 1 MHz = 10^6 Hz, so the frequency in this case is 93.3×10^6 Hz. Also remember that hertz (Hz) is an abbreviation for "cycles per second"; the "cycles" have no units, so the actual units of hertz are simply "per second," equivalent to 1/s. Thus, we find

$$\lambda = \frac{c}{f} = \frac{3 \times 10^8 \frac{\text{m}}{\text{s}}}{93.3 \times 10^6 \frac{1}{\text{s}}} = 3.2 \text{ m}$$

Step 3 Explain: We have found that radio waves with a frequency of 93.3 MHz have a wavelength of 3.2 meters. That is why radio towers are so large—they need to be longer than the size of the waves they are transmitting.

EXAMPLE 2: The wavelength of green visible light (in the middle of the visible spectrum) is about 550 nanometers (1 nm = 10^{-9} m). What is the frequency of this light?

SOLUTION:

Step 1 Understand: Again, we know that visible light obeys the relation $\lambda \times f = c$. This time we are given the wavelength ($\lambda = 550$ nm = 550×10^{-9} m), so we need to solve the equation for the frequency.

Step 2 Solve: We solve the equation ($\lambda \times f = c$) for frequency by dividing both sides by the wavelength (λ), which gives us

$$f = \frac{c}{\lambda}$$

Now we plug in the speed of light (3×10^8 m/s) and the wavelength ($\lambda = 550 \times 10^{-9}$ m) to find

$$f = \frac{c}{\lambda} = \frac{3 \times 10^8 \frac{\text{m}}{\text{s}}}{550 \times 10^{-9}\text{m}} = 5.45 \times 10^{14} \frac{1}{\text{s}}$$

Step 3 Explain: Our answer is about 5.5×10^{14} 1/s; remembering that the units 1/s are called hertz, this answer is equivalent to 5.5×10^{14} Hz, or 550 trillion Hz. Thus, green visible light has a frequency of about 550 trillion Hz. Notice that this frequency is extremely high, which explains in part why it took so long for humans to realize that light behaves like a wave.

EXAMPLE 3: What is the energy of a visible-light photon with wavelength 550 nanometers?

SOLUTION:

Step 1 Understand: The energy of a photon is $E = h \times f$. We are given the photon's wavelength rather than frequency, so we use the result $f = c/\lambda$ from Example 2 to write a formula for calculating energy from wavelength:

$$E = h \times f = h \times \frac{c}{\lambda}$$

Step 2 Solve: We plug in the wavelength ($\lambda = 550 \times 10^{-9}$ m) and Planck's constant ($h = 6.626 \times 10^{-34}$ joule $\times$ s) to find

$$E = h \times \frac{c}{\lambda} = (6.626 \times 10^{-34} \text{ joule} \times \text{s}) \times \frac{3 \times 10^8 \frac{\text{m}}{\text{s}}}{550 \times 10^{-9} \text{ m}} = 3.6 \times 10^{-19} \text{ joule}$$

Step 3 Explain: The energy of a single visible-light photon is about 3.6×10^{-19} joule. Note that this energy for a single photon is extremely small compared to, say, the energy of 100 joules used each second by a 100-watt light bulb.

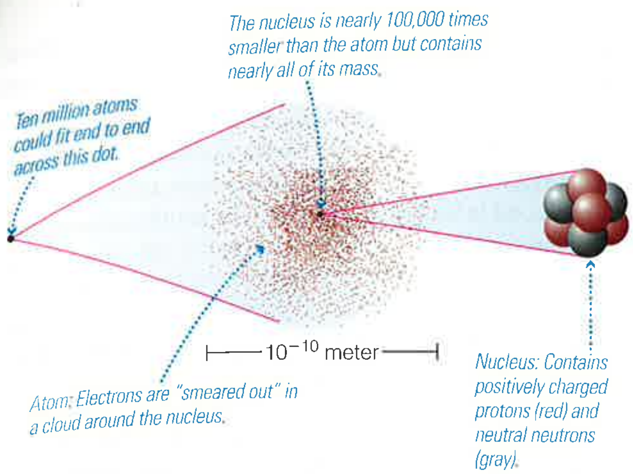

FIGURE 5.8 The structure of a typical atom. Notice that atoms are extremely tiny: The atom shown in the middle is magnified to about 1 billion times its actual size, and the nucleus on the right is magnified to about 100 trillion times its actual size.

center of the atom. The rest of the atom's volume contains electrons, which surround the nucleus. Although the nucleus is very small compared to the atom as a whole, it contains most of the atom's mass, because protons and neutrons are each about 2000 times as massive as an electron. Note that atoms are incredibly small: Millions could fit end to end across the period at the end of this sentence. The number of atoms in a single drop of water (typically, 10^{22} to 10^{23} atoms) may exceed the number of stars in the observable universe.

The properties of an atom depend mainly on the **electrical charge** in its nucleus. Electrical charge is a fundamental physical property that describes how strongly an object will interact with electromagnetic fields; total electrical charge is always conserved, just as energy is always conserved. We define the electrical charge of a proton as the basic unit of positive charge, which we write as +1. An electron has an electrical charge that is precisely opposite that of a proton, so we say it has negative charge (−1). Neutrons are electrically neutral, meaning that they have no charge.

Oppositely charged particles attract and similarly charged particles repel. The attraction between the positively charged protons in the nucleus and the negatively charged electrons that surround it is what holds an atom together. Ordinary atoms have identical numbers of electrons and protons, making them electrically neutral overall.*

Although we can think of electrons as tiny particles, they are not quite like tiny grains of sand and they don't orbit the nucleus the way planets orbit the Sun. Instead, the electrons in an atom form a kind of "smeared out" cloud that surrounds the nucleus and gives the atom its apparent size. The electrons aren't really cloudy, but it is impossible to pinpoint their positions in the atom.

In Figure 5.8, you can see that the electrons give the atom a size far larger than its nucleus even though they represent only a tiny portion of the atom's mass. If you imagine an atom on a scale that makes its nucleus the size of your fist, its electron cloud would be many kilometers wide.

*You may wonder why electrical repulsion doesn't cause the positively charged protons in a nucleus to fly apart from one another. The answer is that an even stronger force, called the *strong force,* overcomes electrical repulsion and holds the nucleus together [Section S4.2].

COMMON MISCONCEPTIONS

The Illusion of Solidity

Bang your hand on a table. Although the table feels solid, it is made almost entirely of empty space! Nearly all the mass of the table is contained in the nuclei of its atoms. But the volume of an atom is more than a trillion times the volume of its nucleus, so the nuclei of adjacent atoms are nowhere near to touching one another. The solidity of the table comes from a combination of electrical interactions between the charged particles in its atoms and the strange quantum laws governing the behavior of electrons. If we could somehow pack all the table's nuclei together, the table's mass would fit into a microscopic speck. Although we cannot pack matter together in this way, nature can and does—in *neutron stars,* which we will study in Chapter 18.

Atomic Terminology You've probably learned the basic terminology of atoms in past science classes, but let's review it just to be sure. Figure 5.9 summarizes the key terminology we will use in this book.

Each different chemical element contains a different number of protons in its nucleus. This number is its **atomic number**. For example, a hydrogen nucleus contains just one proton, so its atomic number is 1. A helium nucleus contains two protons, so its atomic number is 2. The *combined* number of protons and neutrons in an atom is called its **atomic mass number**. The atomic mass number of ordinary hydrogen is

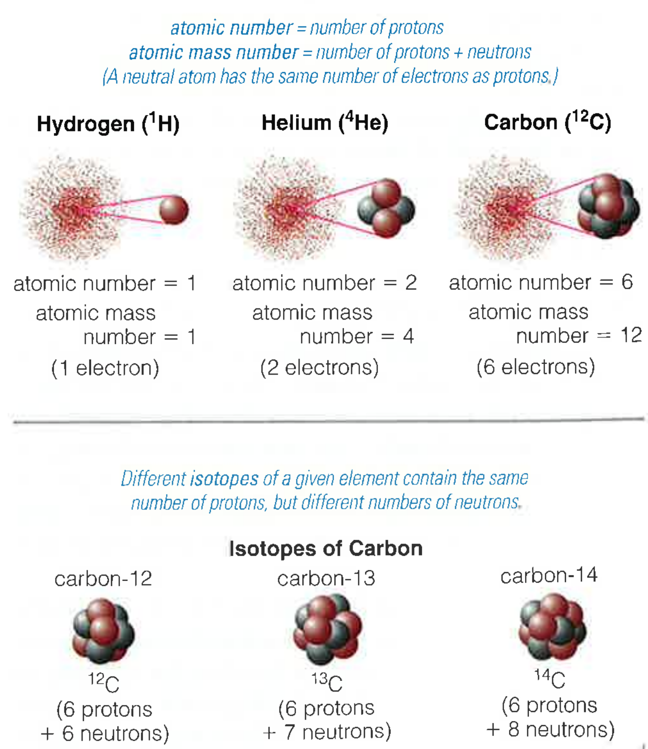

FIGURE 5.9 Terminology of atoms.

1 because its nucleus is just a single proton. Helium usually has two neutrons in addition to its two protons, giving it an atomic mass number of 4. Carbon usually has six protons and six neutrons, giving it an atomic mass number of 12.

Every atom of a given element contains exactly the same number of protons, but the number of neutrons can vary. For example, all carbon atoms have six protons, but they may have six, seven, or eight neutrons. Versions of an element with different numbers of neutrons are called **isotopes** of that element. Isotopes are named by listing their element name and atomic mass number. For example, the most common isotope of carbon has 6 protons and 6 neutrons, giving it atomic mass number $6 + 6 = 12$, so we call it *carbon-12*. The other isotopes of carbon are carbon-13 (six protons and seven neutrons give it atomic mass number 13) and carbon-14 (six protons and eight neutrons give it atomic mass number 14). We can also write the atomic mass number of an isotope as a superscript to the left of the element symbol: ^{12}C, ^{13}C, ^{14}C. We read ^{12}C as "carbon-12."

THINK ABOUT IT

The symbol ^4He represents helium with an atomic mass number of 4. ^4He is the most common form of helium, containing two protons and two neutrons. What does the symbol ^3He represent?

Molecules The number of different material substances is far greater than the number of chemical elements because atoms can combine to form **molecules**. Some molecules consist of two or more atoms of the same element. For example, we breathe O_2, oxygen molecules made of two oxygen atoms. Other molecules, such as water, are made up of atoms of two or more different elements. The symbol H_2O tells us that a water molecule contains two hydrogen atoms and one oxygen atom. Substances composed of molecules with two or more different types of atoms are called **compounds**. Thus, water is a compound.

The chemical properties of a molecule are different from those of its individual atoms. For example, molecular oxygen (O_2) behaves very differently from atomic oxygen (O), and water behaves very differently from pure hydrogen or pure oxygen.

What are the phases of matter?

Interactions between light and matter also depend on the physical state of the matter. Everyday experience tells us that a substance can behave very differently depending on its *phases*, even though it is still made of the same atoms or molecules. For example, molecules of H_2O can exist in three familiar phases: as **solid** ice, as **liquid** water, and as the **gas** we call water vapor. How can the same molecules look and act so differently in these different phases?

You are probably familiar with the idea of a **chemical bond**, the name we give to the interactions between electrons that hold the atoms in a molecule together. For example, we say that chemical bonds hold the hydrogen and oxygen atoms together in a molecule of H_2O. Similar but much weaker interactions among electrons hold together the many water molecules in a block of ice or a pool of water. We can think of the interactions that keep neighboring atoms or molecules close together as other types of bonds.

If we think in terms of bonds, the phases of solid, liquid, and gas differ in the strength of the bonds between neighboring atoms and molecules. Phase changes occur when one type of bond is broken and replaced by another. Changes in either pressure or temperature (or both) can cause phase changes, but it's easier to think first about temperature: As a substance is heated, the average kinetic energy of its particles increases, enabling the particles to break the bonds holding them to their neighbors.

Phase Changes in Water Water is the only familiar substance that we see in all three phases (solid, liquid, gas) in everyday life. The changes in water as it heats up are good examples of the phase changes that occur in all kinds of substances.

At low temperatures, water molecules have a relatively low average kinetic energy. Each molecule is therefore bound tightly to its neighbors, making the *solid* structure of ice. As long as the temperature remains below freezing, the water molecules in ice remain rigidly held together. However, the molecules within this rigid structure are always vibrating, and higher temperature means greater vibrations. If we start with ice at a very low temperature, the molecular vibrations will grow gradually stronger as the temperature rises toward the melting point, which is 0°C at ordinary (sea level) atmospheric pressure.

The melting point is the temperature at which the molecules finally have enough energy to break the solid bonds of ice. The molecules can then move much more freely among one another, allowing the water to flow as a *liquid*. However, the molecules in liquid water are not completely free of one another, as we can tell from the fact that droplets of water can stay intact. Adjacent molecules in liquid water must therefore still be held together by a type of bond, though a much looser bond than the one that holds them together in solid ice.

If we continue to heat the water, the increasing kinetic energy of the molecules will ultimately break the bonds between neighboring molecules altogether. The molecules will then be able to move freely, and freely moving particles constitute what we call a *gas*. Above the boiling point (100°C at sea level), all the bonds between adjacent molecules are broken so that the water can exist only as a gas.

We see ice melting into liquid water and liquid water boiling into gas so often that it's tempting to think that's the end of the story. However, a little thought should convince you that the reality has to be more complex. For example, you know that Earth's atmosphere contains water vapor that condenses to form clouds and rain. But Earth's surface temperature is well below the boiling point of water, so how is it that our atmosphere can contain water in the gas phase?

You'll understand the answer if you remember that temperature is a measure of the *average* kinetic energy of the particles in a substance [Section 4.3]; individual particles may have substantially lower or higher energies than the average. Thus, even at the low temperatures at which most water molecules are bound together as ice or liquid, a few molecules will always have enough energy to break free of their neighbors and enter the gas phase. In other words, some gas (water vapor) is always present along with solid ice or liquid water. The process by which molecules escape from a solid is called **sublimation**, and the process by which molecules escape from a liquid is called **evaporation.** Higher temperatures lead to higher rates of sublimation or evaporation.

THINK ABOUT IT

Global warming is expected to increase Earth's average temperature by up to a few degrees over the coming century. Based on what you've learned about phase changes, how would you expect global warming to affect the total amount of cloud cover on Earth? Explain.

Molecular Dissociation and Ionization Above the boiling point, all the water will have entered the gas phase. What happens if we continue to raise the temperature?

The molecules in a gas move freely, but they often collide with one another. As the temperature rises, the molecules move faster and the collisions become more violent. At high enough temperatures, the collisions will be so violent that they can break the chemical bonds holding individual water molecules together. The molecules then split into pieces, a process we call **molecular dissociation.** (In the case of water, molecular dissociation usually frees one hydrogen atom and leaves a negatively charged molecule that consists of one hydrogen atom and one oxygen atom [OH]; at even higher temperatures, the OH dissociates into individual atoms.)

At still higher temperatures, collisions can break the bonds holding electrons around the nuclei of individual atoms, allowing the electrons to go free. The loss of one or more negatively charged electrons leaves the remaining atom with a net positive charge. Such charged atoms are called **ions**. The process of stripping electrons from atoms is called **ionization.**

Thus, at temperatures of several thousand degrees, what once was water becomes a hot gas consisting of freely moving electrons and positively charged ions of hydrogen and oxygen. This type of hot gas, in which atoms have become ionized, is called a **plasma.** Because a plasma contains many charged particles, its interactions with light are different from those of a gas consisting of neutral atoms, which is one reason that plasma is sometimes referred to as "the fourth phase of matter." However, because the electrons and ions are not bound to one another, it is also legitimate to call plasma a gas. That is why we sometimes say that the Sun is made of hot gas and sometimes say that it is made of plasma; both statements are correct.

The degree of ionization in a plasma depends on its temperature and composition. A neutral hydrogen atom contains only one electron, so hydrogen can be ionized only once; the remaining hydrogen ion, designated H^+, is simply a proton. Oxygen, with atomic number 8, has eight electrons when it is neutral, so it can be ionized multiple times. *Singly ionized* oxygen is missing one electron, so it has a charge of +1 and is designated O^+. *Doubly ionized* oxygen, or O^{++}, is missing two electrons; *triply ionized* oxygen, or O^{+3}, is missing three

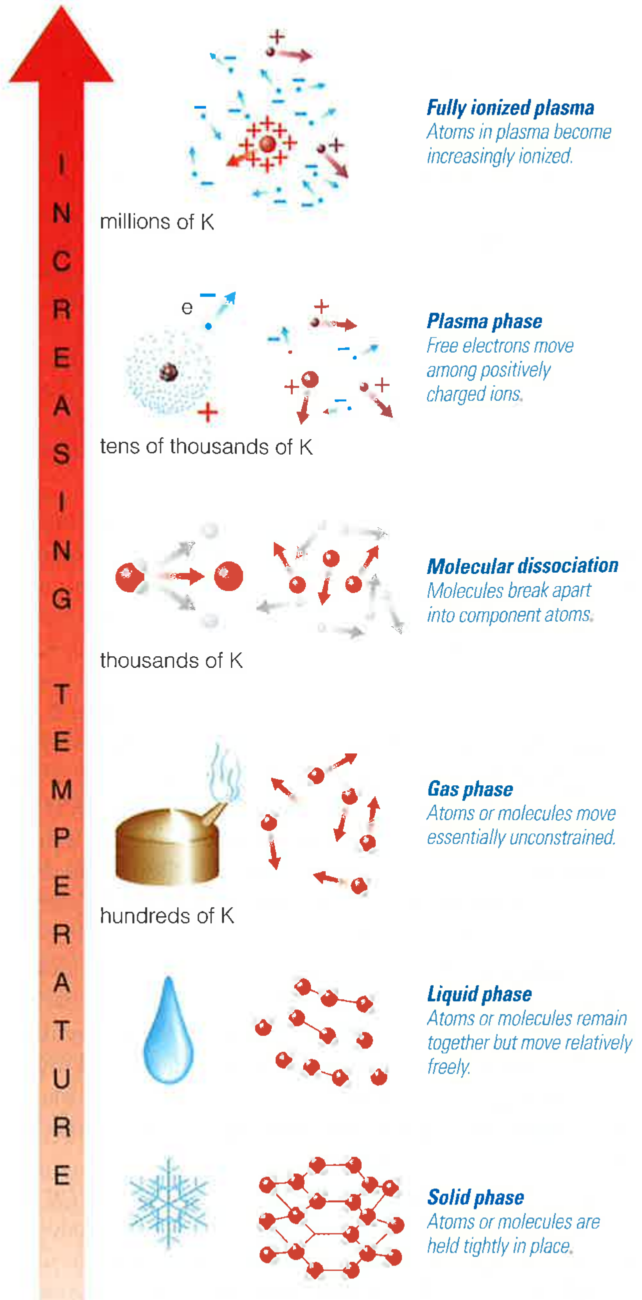

FIGURE 5.10 The general progression of phase changes in water.

electrons; and so on. At temperatures of several million degrees, oxygen can be *fully ionized,* in which case all eight electrons are stripped away and the remaining ion has a charge of +8.

Figure 5.10 summarizes the changes that occur as we heat water from ice to a fully ionized plasma. Other chemical substances go through similar phase changes, but the changes generally occur at different temperatures for different substances.

Phases and Pressure Temperature is the primary factor determining the phase of a substance and the ways in which light interacts with it, but pressure also plays a role. You're undoubtedly familiar with the idea of pressure in an everyday sense: For example, you can put more pressure on your arm

COMMON MISCONCEPTIONS

One Phase at a Time?

In daily life, we usually think of H_2O as being in just one phase (solid ice, liquid water, or water vapor) at one time, with the phase depending on the temperature. In reality, two or even all three phases can exist at the same time. In particular, some sublimation *always* occurs over solid ice, and some evaporation *always* occurs over liquid water. You can tell that evaporation always occurs, because an uncovered glass of water will gradually empty as the liquid evaporates into gas. You can see that sublimation occurs by watching the snow pack after a winter storm: Even if the snow doesn't melt into liquid, it will gradually disappear as the ice sublimates into water vapor.

by squeezing it. In science, we use a more precise definition: **pressure** is the *force per unit area* pushing on an object's surface. You feel more pressure when you squeeze your arm because squeezing increases the force on each square centimeter of your arm's surface. Similarly, piling rocks on a table increases the weight (force) on the table, which therefore increases the pressure on the surface of the table; if the pressure becomes too great, the table breaks. The gas in an atmosphere also creates pressure, because the weight of the gas bears down on everything beneath it. For example, at sea level on Earth, the weight of the atmosphere creates a pressure of about 14.7 pounds per square inch. That is, the total weight of all the air above each square inch of Earth's surface is about 14.7 pounds [Section 10.1].

Pressure can affect phases in a variety of ways. For example, deep inside Earth, the pressure is so high that Earth's inner metal core remains solid, even though the temperature is high enough that the metal would melt into liquid under less extreme pressure conditions [Section 9.1]. On a planetary surface, atmospheric pressure can determine whether water is stable in liquid form.

Remember that liquid water is always evaporating (or ice sublimating) at a low level, because a few molecules randomly get enough energy to break the bonds holding them to their neighbors. On Earth, enough liquid water has evaporated from the oceans to make water vapor an important ingredient of our atmosphere. Some of these atmospheric water vapor molecules collide with the ocean surface, where they can "stick" and rejoin the ocean—essentially the opposite of evaporation (Figure 5.11). The greater the pressure created by water vapor molecules in our atmosphere,* the higher the rate at which water molecules return to the ocean. This direct return of water vapor molecules from the atmosphere helps keep the total amount of water in Earth's oceans fairly stable.† On the Moon, where the lack of atmosphere means no pressure from water vapor at all, liquid water would evaporate quite quickly (as long as the temperature were high enough that it did not freeze first). The same is true on Mars, because the atmosphere lacks enough water vapor to balance the rate of evaporation.

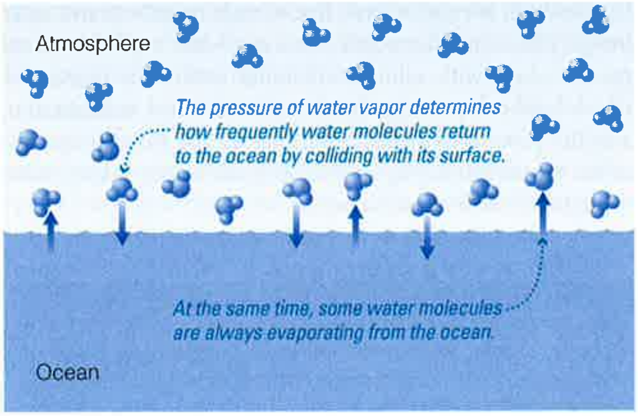

FIGURE 5.11 Evaporation of water molecules from the ocean is balanced in part by molecules of water vapor in Earth's atmosphere returning to the ocean. The rate at which these molecules return is directly related to the pressure created by water vapor in the atmosphere.

High pressure can also cause gases to dissolve in liquid water. For example, sodas are made by putting water in contact with high-pressure carbon dioxide gas. Because of the high pressure, many more carbon dioxide molecules enter the water than are released, so the water becomes "carbonated"—that is, it has a lot of dissolved carbon dioxide. When you open a bottle of carbonated water, exposing it to air with ordinary pressure, the dissolved carbon dioxide quickly bubbles up and escapes.

How is energy stored in atoms?

Now that we have reviewed the structure of matter and how its phase depends on temperature and pressure, it is time to return to the primary goal of this chapter: understanding how we learn about distant objects by studying their light. To produce light, these objects must somehow transform energy contained in matter into the vibrations of electric and magnetic fields that we call light. We therefore need to focus on the charged particles within atoms, particularly the electrons, because only particles that have charge can interact with light.

Atoms actually contain energy in three different ways. First, by virtue of their mass, they possess mass-energy in the amount mc^2. Second, they possess kinetic energy by virtue of their motion. Third, and most important to reading the messages encoded in light, atoms contain *electrical potential energy* that depends on the arrangement of their electrons around their nuclei. To interpret the messages carried by light, we must understand how electrons store and release their electrical potential energy.

Energy Levels in Atoms The energy stored by electrons in atoms has a strange but important property: The electrons can have only particular amounts of energy, and not other energies in between. As an analogy, suppose you're washing

*Technically, this is known as the *vapor pressure* of water in the atmosphere. We can also measure vapor pressure for other atmospheric constituents, and the total gas pressure is the sum of all the individual vapor pressures.

†Rain and snow also contribute, of course; however, even if Earth's temperature rose enough so that raindrops and snowflakes could no longer form, only a small fraction of Earth's ocean water would evaporate before the return rate of water vapor molecules balanced the evaporation rate.

windows on a building. If you use an adjustable platform to reach high windows, you can stop the platform at any height above the ground. But if you use a ladder, you can stand only at *particular* heights—the heights of the rungs of the ladder—and not at any height in between. The possible energies of electrons in atoms are like the possible heights on a ladder. Only a few particular energies are possible, and energies between these special few are not possible. The possible energies are known as the **energy levels** of an atom.

Figure 5.12 shows the energy levels of hydrogen, the simplest of all elements. The energy levels are labeled on the left in numerical order and on the right with energies in units of *electron-volts*, or *eV* for short (1 eV $= 1.60 \times 10^{-19}$ joule). The lowest possible energy level—called level 1 or the *ground state*—is defined as an energy of 0 eV. Each of the higher energy levels (sometimes called *excited states*) is labeled with the extra energy of an electron in that level compared to the ground state.

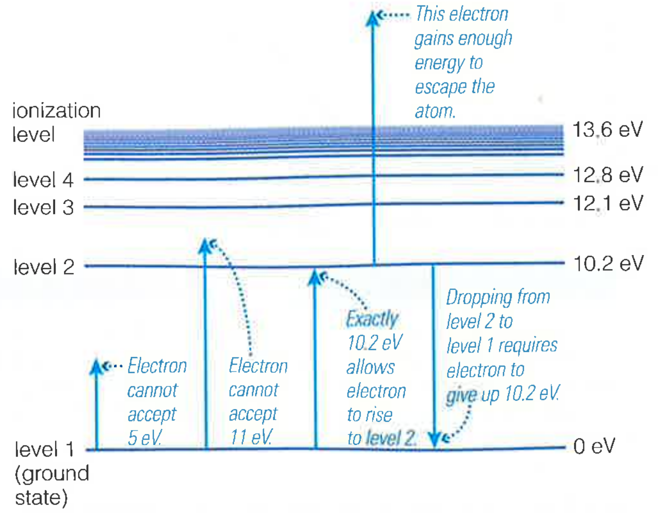

FIGURE 5.12 Energy levels for the electron in a hydrogen atom. The electron can change energy levels only if it gains or loses the amount of energy separating the levels. If the electron gains enough energy to reach the ionization level, it can escape from the atom, leaving behind a positively charged ion. (The many levels between level 4 and the ionization level are not labeled.)

Energy Level Transitions An electron can rise from a low energy level to a higher one or fall from a high level to a lower one; such changes are called **energy level transitions.** Because energy must be conserved, energy level transitions can occur only when an electron gains or loses the specific amount of energy separating two levels. For example, an electron in level 1 can rise to level 2 only if it gains 10.2 eV of energy. If you try to give the electron 5 eV of energy, it won't accept it because that is not enough energy to reach level 2. Similarly, if you try to give it 11 eV, it won't accept it because it is too much for level 2 but not enough to reach level 3. Once in level 2, the electron can return to level 1 by giving up 10.2 eV of energy. Figure 5.12 shows several examples of allowed and disallowed energy level transitions.

Notice that the amount of energy separating the various levels gets smaller at higher levels. For example, it takes more energy to raise the electron from level 1 to level 2 than from level 2 to level 3, which in turn takes more energy than the transition from level 3 to level 4. If the electron gains enough energy to reach the *ionization level,* it escapes the atom completely, thereby ionizing the atom. Any excess energy beyond the amount needed for ionization becomes kinetic energy of the free-moving electron.

THINK ABOUT IT

Are there any circumstances under which an electron in a hydrogen atom can lose 2.6 eV of energy? Explain.

Quantum Physics If you think about it, the idea that electrons in atoms are restricted to particular energy levels is quite bizarre. It is as if you had a car that could go around a track only at particular speeds and not at speeds in between. How strange it would seem if your car suddenly changed its speed from 5 miles per hour to 20 miles per hour without first passing through a speed of 10 miles per hour! In scientific terminology, the electron's energy levels in an atom are said to be *quantized,* and the study of the energy levels of electrons (and other particles) is called *quantum physics* (or *quantum mechanics*). We will explore some of the astonishing implications of quantum physics in Chapter S4.

Electrons have quantized energy levels in all atoms, not just in hydrogen. Moreover, the allowed energy levels differ from element to element and from one ion of an element to another ion of the same element. Even molecules have quantized energy levels. As we will see shortly, the different energy levels of different atoms and molecules allow light to carry "fingerprints" that can tell us the chemical composition of distant objects.

5.4 LEARNING FROM LIGHT

Matter leaves its fingerprints whenever it interacts with light. Examining the color of an object is a crude way of studying the clues left by the matter it contains. For example, a red shirt absorbs all visible photons except those in the red part of the spectrum, so we know that it must contain a dye with these special light-absorbing characteristics. If we take light and disperse it into a spectrum, we can see the spectral fingerprints more clearly. For example, the photograph that opens this chapter (p. 142) shows the Sun's visible-light spectrum in great detail, with the rainbow of color stretching in horizontal rows from the upper left to the lower right of the photograph. We see similar dark or bright lines when we look at almost any spectrum, whether it is the spectrum of the flame from the gas grill in someone's backyard or the spectrum of a distant galaxy whose light we collect with a gigantic telescope. As long as we collect enough light to see details in the spectrum, we can learn many fundamental properties of the object we are viewing, no matter how far away the object is located.

The process of obtaining a spectrum and reading the information it contains is called **spectroscopy**. Spectra often look much like rainbows, at least if they are made only from