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.

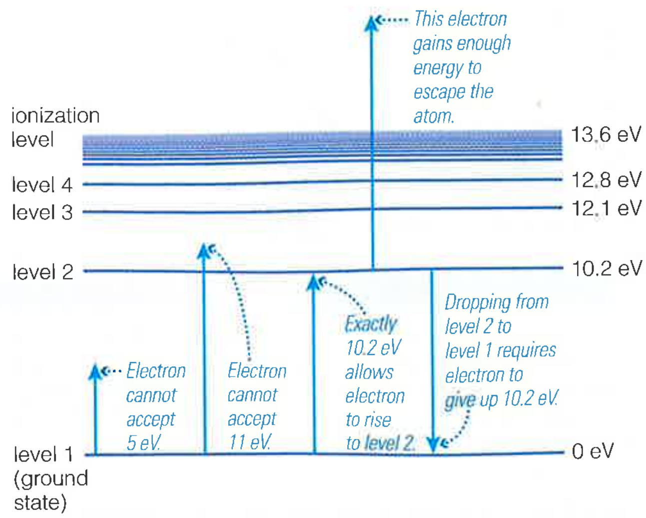



**FIGURE 5.12** Energy levels for the electron in a hydrogen atom. The electron can change energy levels only if it gains or loses the amount of energy separating the levels. If the electron gains enough energy to reach the ionization level, it can escape from the atom, leaving behind a positively charged 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

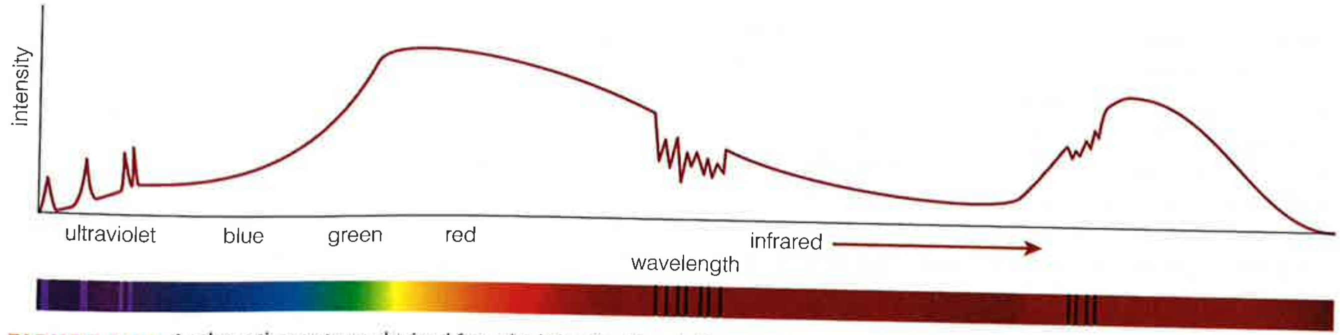



**FIGURE 5.13** A schematic spectrum obtained from the light of a distant object. The "rainbow" at bottom shows how the light would appear when viewed through a prism or diffraction grating; of course, our eyes cannot see the ultraviolet or infrared light. The graph shows the corresponding intensity of the light at each wavelength. The intensity is high where the rainbow is bright and low where it is dim (such as in places where the rainbow shows dark lines).

visible light. However, for the purposes of examining spectra in greater detail, it's often more useful to display them as graphs that show the amount of radiation, or **intensity**, at different wavelengths. The graphs of intensity in this book simply show the relative amount of energy received from an object in each wavelength of light. In other words, the intensity at a given wavelength is proportional to the number of photons observed at that wavelength times the energy of those photons.

For example, Figure 5.13 shows a schematic spectrum of light from a celestial body such as a planet. Notice that this spectrum extends all the way from the ultraviolet on the left to the infrared on the right, which means the graph is showing us details that would be invisible to our eyes in a photograph. At wavelengths where a lot of light is coming from the celestial body, the intensity is high, while at wavelengths where there is little light, the intensity is low.

Our goal in this section is to learn how to interpret astronomical spectra. The bumps and wiggles in Figure 5.13 arise from several different processes, making it a good case study for spectral interpretation. We'll consider these processes one at a time, then return to interpret Figure 5.13 at the end of this section.

MA™ **Light and Spectroscopy Tutorial, Lessons 2–4**

## What are the three basic types of spectra?

Let's begin our study of spectra by learning how to classify them. Laboratory studies show that spectra come in three basic types: continuous, emission line, and absorption line. Figure 5.14 shows an example of each type, along with the conditions that produce it. (The rules that specify these conditions are often called *Kirchhoff's laws.*) Study the figure carefully to notice the following key ideas:

1. The spectrum of an ordinary (incandescent) light bulb is a rainbow of color. Because the rainbow spans a broad range of wavelengths without interruption, we call it a **continuous spectrum**.
2. A thin or low-density cloud of gas does not produce a continuous spectrum. Instead, it emits light only at specific wavelengths that depend on its composition and temperature. The spectrum therefore consists of bright **emission lines** against a black background, and is called an **emission line spectrum**.
3. If the cloud of gas lies between us and a light bulb, we still see most of the continuous light emitted by the light bulb. However, the cloud absorbs light of specific wavelengths, so the spectrum shows dark **absorption lines** over the background rainbow from the light bulb.* We call this an **absorption line spectrum**.

Each of the three spectra is shown both as a band of light (essentially as it would be recorded on a photograph) and as a graph of intensity versus wavelength. Notice that absorption lines appear as dips on a background of relatively high-intensity light, while emission lines look like spikes on a background with little or no intensity.

We can apply some of these ideas to the solar spectrum that opens this chapter. Notice that it shows numerous absorption lines over a background rainbow of color. This tells us that we are essentially looking at a hot light source through gas that is absorbing some of the colors, much as we see when looking through the cloud of gas to the light bulb in Figure 5.14c. For the solar spectrum, the hot light source is the hot interior of the Sun, while the "cloud" is the relatively cool and low-density layer of gas that makes up the Sun's visible surface, or *photosphere* [Section 14.1].

## How does light tell us what things are made of?

We have just seen *how* different viewing conditions lead us to see different types of spectra, so we are now ready to discuss *why.* Let's start with emission and absorption line spectra. Emission and absorption lines form as a direct consequence of the fact that each type of atom, ion, or molecule possesses a unique set of energy levels. This fact allows us to learn the compositions of distant objects in the universe, as we can see

*More technically, we'll see an absorption line spectrum as long as the cloud is cooler in temperature than the source of background light, which in this case means the light bulb filament.

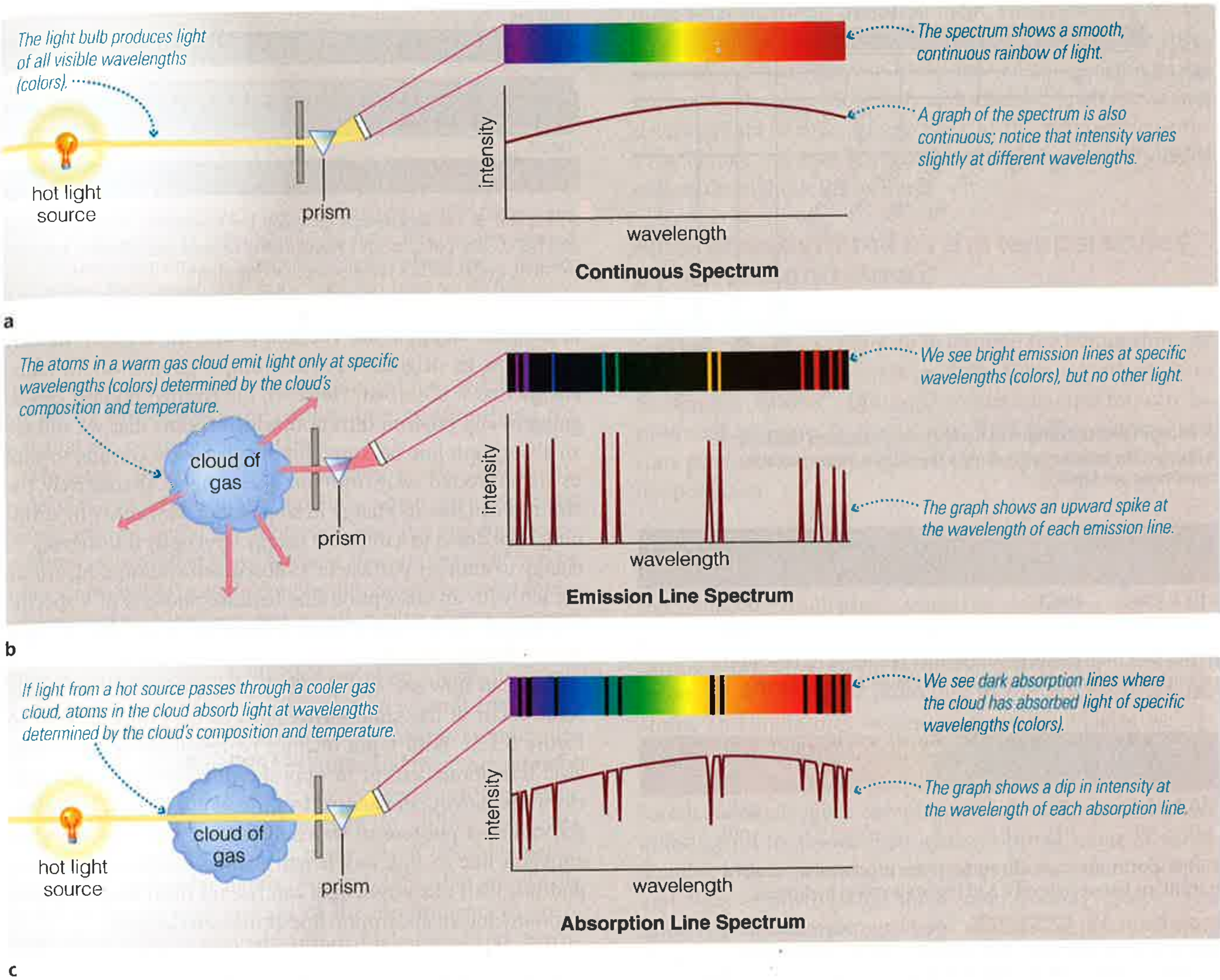



**FIGURE 5.14** Interactive figure These diagrams show examples of the conditions under which we see the three basic types of spectra.

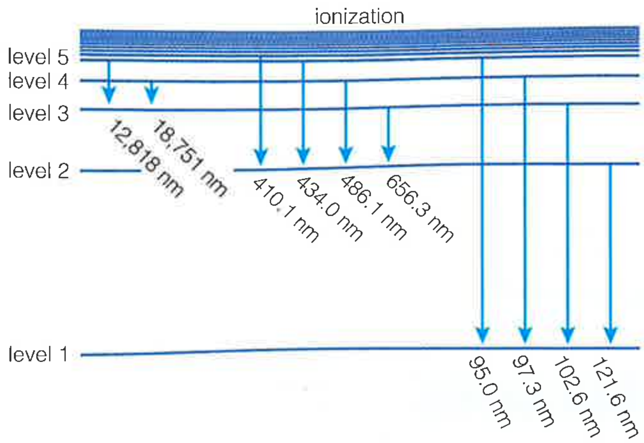
by considering what happens in a cloud of gas consisting solely of hydrogen atoms.

**Emission Line Spectra** The atoms in any cloud of gas are constantly colliding with one another, exchanging energy in each collision. Most of the collisions simply send the atoms flying off in new directions. However, a few of the collisions transfer the right amount of energy to bump an electron from a low energy level to a higher energy level.

Electrons can't stay in higher energy levels for long. They always fall back down to the ground state, level 1, usually in a tiny fraction of a second. The energy the electron loses when it falls to a lower energy level must go somewhere, and often it goes to *emitting* a photon of light. The emitted photon must have the same amount of energy that the electron loses, which means that it has a specific wavelength (and frequency). Figure 5.15a shows the energy levels in hydrogen that we saw in Figure 5.12, but it is also labeled with the wavelengths of the photons emitted by various downward transitions of an electron from a higher energy level to a lower one. For example, the transition from level 2 to level 1 emits an ultraviolet photon of wavelength 121.6 nm, and the transition from level 3 to level 2 emits a red visible-light photon of wavelength 656.3 nm.*

Although electrons that rise to higher energy levels in a gas quickly return to level 1, new collisions can raise other electrons into higher levels. As long as the gas remains moderately warm, collisions are always bumping some electrons into levels from which they fall back down and emit photons with some of the wavelengths shown in Figure 5.15a. The gas therefore emits light with these specific wavelengths. That is why a warm gas cloud produces an emission line spectrum, as shown in Figure 5.15b. The bright emission lines appear at the wavelengths that correspond to downward transitions of electrons, and the rest of the spectrum is dark (black). The specific set of lines that we see depends on the cloud's temperature as well as its composition: At higher temperatures, electrons are more likely to be bumped to higher energy levels.

*Astronomers call transitions between level 1 and other levels the *Lyman* series of transitions. The transition between level 1 and level 2 is Lyman $\alpha$, between level 1 and level 3 Lyman $\beta$, and so on. Similarly, transitions between level 2 and higher levels are called *Balmer* transitions. Other sets of transitions also have names.




**a** Energy level transitions in hydrogen correspond to photons with specific wavelengths. Only a few of the many possible transitions are labeled.


**b** This spectrum shows emission lines produced by downward transitions between higher levels and level 2 in hydrogen.


**c** This spectrum shows absorption lines produced by upward transitions between level 2 and higher levels in hydrogen.

**FIGURE 5.15** Interactive Figure An atom emits or absorbs light only at specific wavelengths that correspond to changes in the atom's energy as an electron undergoes transitions between its allowed energy levels.

**THINK ABOUT IT**

If nothing continues to heat the hydrogen gas, all the electrons eventually will end up in the lowest energy level (the ground state, or level 1). Use this fact to explain why we should *not* expect to see an emission line spectrum from a very cold cloud of hydrogen gas.

**Absorption Line Spectra** Now, suppose a light bulb illuminates the hydrogen gas from behind (as in Figure 5.14c). The light bulb emits light of all wavelengths, producing a spectrum that looks like a rainbow of color. However, the hydrogen atoms can absorb those photons that have the right amount of energy needed to raise an electron from a low energy level to a higher one. Figure 5.15c shows the result. It is an absorption line spectrum, because the light bulb produces a continuous rainbow of color while the hydrogen atoms absorb light at specific wavelengths.

Given that electrons in high energy levels quickly return to lower levels, you might wonder why photons emitted in downward transitions don't cancel out the effects of those absorbed in upward transitions. Finding the answer requires looking more deeply at what happens to the absorbed photons. Two things can happen after an electron absorbs a photon and rises to a higher energy level. The first is that the electron quickly returns to its original level, emitting a photon of the same energy that it absorbed. However, the emitted photon can be going in any random direction, which means that we still see an absorption line because photons that were coming toward us are redirected away from our line of sight. Alternatively, the electron can lose its energy in some other way, either by dropping back down to a different energy level or by transferring its energy to another particle in a subsequent collision. Again, we are left with an absorption line because photons of a specific wavelength have been removed from the spectrum of light that's coming toward us.

You can now see why the dark absorption lines in Figure 5.15c occur at the same wavelengths as the emission lines in Figure 5.15b: Both types of lines represent the same energy level transitions, except in opposite directions. For example, electrons moving downward from level 3 to level 2 in hydrogen can emit photons of wavelength 656.3 nm (producing an emission line at this wavelength), while electrons absorbing photons with this wavelength can rise up from level 2 to level 3 (producing an absorption line at this wavelength).


**FIGURE 5.16** Visible-light emission line spectra for helium, sodium, and neon. The patterns and wavelengths of lines are different for each element, giving each a unique spectral fingerprint.

**Chemical Fingerprints** The fact that hydrogen emits and absorbs light at specific wavelengths makes it possible to detect its presence in distant objects. For example, imagine that you look through a telescope at an interstellar gas cloud, and its spectrum looks like that shown in Figure 5.15b. Because only hydrogen produces this particular set of lines, you can conclude that the cloud is made of hydrogen. In essence, the spectrum contains a "fingerprint" left by hydrogen atoms.

Real interstellar clouds are not made solely of hydrogen. However, the other chemical constituents in the cloud leave fingerprints on the spectrum in much the same way. Every type of atom has its own unique spectral fingerprint, because it has its own unique set of energy levels. For example, Figure 5.16 shows emission line spectra for helium, sodium, and neon.

Not only does each chemical element produce a unique spectral fingerprint, but *ions* of a particular element (atoms that are missing one or more electrons) produce fingerprints different from those of neutral atoms (Figure 5.17). For example, the spectrum of doubly ionized neon ($Ne^{++}$) is different from that of singly ionized neon ($Ne^{+}$), which in turn is different from that of neutral neon (Ne). These differences can help us determine the temperature of a hot gas or plasma, and they come in handy when we're trying to measure the surface temperatures of stars [Section 15.1]. At

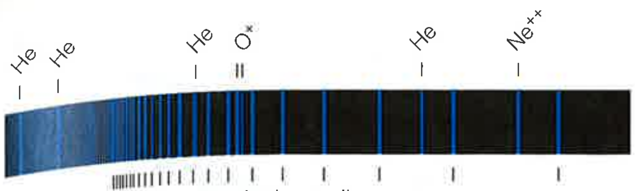
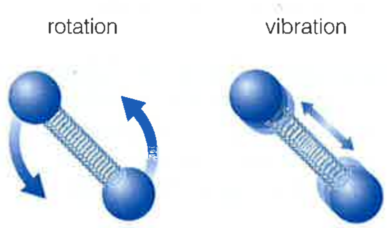



**FIGURE 5.17** The emission line spectrum of the Orion Nebula in a portion of the ultraviolet (about 350–400 nm). The lines are labeled with the chemical elements or ions that produce them (He = helium; O = oxygen; Ne = neon). The many hydrogen lines are all transitions from high levels to level 2.

higher temperatures, more highly charged ions will be present, so we can estimate the temperature by identifying the ions that are creating spectral lines.

Molecules also produce spectral fingerprints. Like atoms, molecules can produce spectral lines when their electrons change energy levels. But molecules can also produce spectral lines in two other ways. Because they are made of two or more atoms bound together, molecules can vibrate and rotate (Figure 5.18a). Vibration and rotation also require energy, and the possible energies of rotation and vibration in molecules are quantized much like electron energy levels in atoms. A molecule can absorb or emit a photon when it changes its rate of vibration or rotation. The energy changes in molecules are usually smaller than those in atoms and therefore produce lower-energy photons, and the energy levels also tend to be bunched more closely together than in atoms. Molecules therefore produce spectra with many sets of tightly bunched lines, called **molecular bands** (Figure 5.18b), that are usually found in the infrared portion of the electromagnetic spectrum. That is one reason why infrared telescopes and instruments are so important to astronomers.

Over the past century, scientists have conducted laboratory experiments to identify the spectral lines of every chemical element and of many ions and molecules. Thus, when we see lines in the spectrum of a distant object, we can usually determine what chemicals produced them. For example, if we see spectral lines of hydrogen, helium, and carbon in the spectrum of a distant star, we know that all three elements are present in the star. Moreover, with detailed analysis, we can determine the relative proportions of the various elements. That is how we have learned the chemical compositions of objects throughout the universe.




**a** We can think of a two-atom molecule as two balls connected by a spring. Although this model is simplistic, it illustrates how molecules can rotate and vibrate. The rotations and vibrations can have only particular amounts of energy and therefore produce unique spectral fingerprints.

**b** This spectrum of molecular hydrogen ($H_2$) consists of lines bunched into broad molecular bands.

**FIGURE 5.18** Like atoms and ions, molecules also emit or absorb light at specific wavelengths.

## How does light tell us the temperatures of planets and stars?

We have seen how emission and absorption line spectra form and how we can use them to determine the composition of a cloud of gas. Now we are ready to turn our attention to continuous spectra. Although continuous spectra can be produced in more than one way, light bulbs, planets, and stars produce them in a way that can help us determine their temperatures.

**Thermal Radiation: Every Body Does It** In a cloud of gas that produces a simple emission or absorption line spectrum, the individual atoms or molecules are essentially independent of one another. Most photons pass easily through such a gas, except those that cause energy level transitions in the atoms or molecules of the gas. However, the atoms and molecules within most of the objects we encounter in everyday life—such as rocks, light bulb filaments, and people—cannot be considered independent and therefore have much more complex sets of energy levels. These objects tend to absorb light across a broad range of wavelengths, which means light cannot easily pass through them and light emitted inside them cannot easily escape. The same is true of almost any large or dense object, including planets and stars.

In order to understand the spectra of such objects, let's consider an idealized case in which an object absorbs all photons that strike it and does not allow photons inside it to escape easily. Photons trying to escape are quickly absorbed by an atom or molecule, which quickly reemits the photon—but often with a slightly different wavelength and in a different direction.* In effect, the emitted photons bounce randomly around inside the object, constantly exchanging energy with the object's atoms or molecules. By the time the photons finally escape the object, their radiative energies have become randomized so that they are spread over a wide range of wavelengths. The wide wavelength range of the photons explains why the spectrum of light from such an object is smooth, or *continuous*, like a pure rainbow without any absorption or emission lines.

Most important, the spectrum from such an object depends on only one thing: the object's *temperature*. To understand why, remember that temperature represents the average kinetic energy of the atoms or molecules in an object [Section 4.3]. Because the randomly bouncing

*One of the reasons that photons are reemitted with a slightly different wavelength has to do with the Doppler effect (see Section 5.5). If the absorbing atom (or molecule) is moving, there will be a shift in wavelength that depends on the atom's velocity and the direction in which the photon escapes.

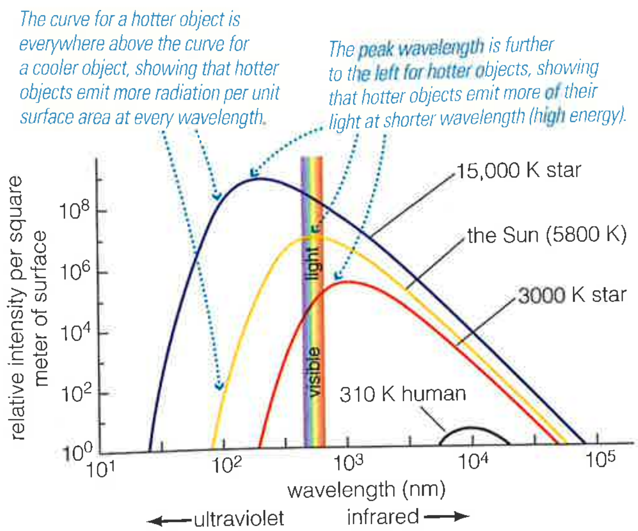



**FIGURE 5.19** Interactive Figure Graphs of idealized thermal radiation spectra demonstrate the two laws of thermal radiation: (1) Each square meter of a hotter object's surface emits more light at all wavelengths; (2) hotter objects emit photons with a higher average energy. Notice that the graph uses power-of-10 scales on both axes, so that we can see all the curves even though the differences between them are quite large.

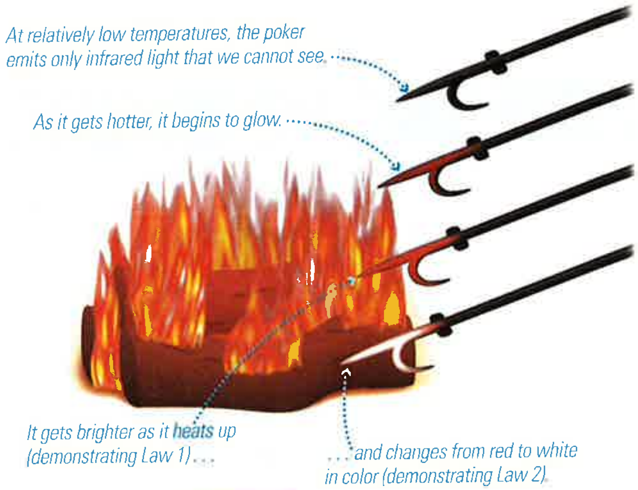



**FIGURE 5.20** Interactive Figure A fireplace poker shows the two laws of thermal radiation in action.

photons interact so many times with those atoms or molecules, they end up with energies that match the kinetic energies of the object's atoms or molecules—which means the photon energies depend only on the object's temperature, regardless of what the object is made of. The temperature dependence of this light explains why we call it **thermal radiation** (sometimes known as *blackbody radiation*), and why its spectrum is called a **thermal radiation spectrum.** Thermal radiation spectra are the most common type of continuous spectra.

No real object emits a perfect thermal radiation spectrum, but almost all familiar objects—including the Sun, the planets, rocks, and even you—emit light that approximates thermal radiation. Figure 5.19 shows graphs of the idealized thermal radiation spectra of three stars and a human, each with its temperature given on the Kelvin scale (see Figure 4.12). Be sure to notice that these spectra show the intensity of light *per unit surface area,* not the total amount of light emitted by the object. For example, a very large 3000 K star can emit more total light than a small 15,000 K star, even though the hotter star emits much more light per unit area.

**The Two Laws of Thermal Radiation** If you compare the spectra in Figure 5.19, you'll see that they obey two laws of thermal radiation:

- Law 1 (Stefan-Boltzmann law): *Each square meter of a hotter object's surface emits more light at all wavelengths.* For example, each square meter on the surface of the 15,000 K star emits a lot more light at every wavelength than each square meter of the 3000 K star, and the hotter star emits light at some ultraviolet wavelengths that the cooler star does not emit at all.
- Law 2 (Wien's law ["Wien" is pronounced *veen*]): *Hotter objects emit photons with a higher average energy,* which means a shorter average wavelength. That is why the peaks of the spectra are at shorter wavelengths for hotter objects. For example, the peak for the 15,000 K star is in ultraviolet light, the peak for the 5800 K Sun is in visible light, and the peak for the 3000 K star is in the infrared.

You can see these laws in action with a fireplace poker (Figure 5.20). While the poker is still relatively cool, it emits only infrared light, which we cannot see. As it gets hot (above about 1500 K), it begins to glow with visible light, and it glows more brightly as it gets hotter, demonstrating the first law. Its color demonstrates the second law. At first it glows "red hot," because red light has the longest wavelengths of visible light. As it gets even hotter, the average wavelength of the emitted photons moves toward the blue (short wavelength) end of the visible spectrum. The mix of colors emitted at this higher temperature makes the poker look white to your eyes, which is why "white hot" is hotter than "red hot."

**SEE IT FOR YOURSELF**

Find a light that has a dimmer switch. What happens to the bulb temperature (which you can check by placing your hand near it) as you turn the switch up? How does the light change color? Explain how these observations demonstrate the two laws of thermal radiation.

Because thermal radiation spectra depend only on temperature, we can use them to measure the temperatures of distant objects. In many cases we can estimate temperatures simply from the object's color. Notice that while hotter objects emit more light at *all* wavelengths, the biggest difference appears at the shortest wavelengths. At human body temperature of about 310 K, people emit mostly in the infrared and emit no visible light at all—which explains why we don't glow in the dark! A relatively cool star, with a 3000 K surface temperature, emits mostly red light. That is why some bright stars in our sky, such as Betelgeuse (in Orion) and Antares (in Scorpius), appear reddish in color. The Sun's 5800 K surface emits most strongly in green light (around 500 nm), but the Sun looks yellow or white to our eyes because it also emits other colors throughout the visible spectrum. Hotter stars emit mostly in the ultraviolet but appear blue-white in color because our eyes cannot see their ultraviolet light. If an object were heated to a temperature of millions of degrees, it would radiate mostly X rays. Some astronomical objects are indeed hot enough to emit X rays, such as disks of gas encircling exotic objects like neutron stars and black holes (see Chapter 18).

## How do we interpret an actual spectrum?

The spectra of real astronomical objects are usually combinations of the three idealized types of spectra (continuous or thermal, absorption line, and emission line). They may also show features produced by reflection or scattering.

**Reflected Light Spectra** An object that reflects light will have the spectrum of the light shining on it, minus the light it absorbs. For example, a red sweatshirt absorbs blue light and reflects red light, so its visible spectrum looks like the spectrum of sunlight (or lamp light) but with blue light missing. In the same way that we can distinguish lemons from limes, we can use color information in reflected light to learn about celestial objects. Different fruits, different rocks, and even different atmospheric gases reflect and absorb light at different wavelengths. Although the absorption features that show up in spectra of reflected light are not as distinct as the emission and absorption lines for thin gases, they still provide useful information. For example, the surface materials of a planet determine how much light of different colors is reflected or absorbed. The reflected light gives the planet its color, while the absorbed light heats the surface and helps determine its temperature.

**Putting It All Together** Figure 5.21 shows the same spectrum we began with in Figure 5.13, but this time with labels indicating the processes responsible for its various features. The thermal emission peaks in the infrared, showing a surface temperature of about 225 K, well below the 273 K freezing point of water. The absorption bands in the infrared come mainly from carbon dioxide, indicating a carbon dioxide atmosphere. The emission lines in the ultraviolet come from hot gas in a high, thin layer of the object's atmosphere. The reflected light looks like the Sun's 5800 K thermal radiation except that much of the blue light is missing, so the object must be reflecting sunlight and must look red in color. Perhaps by now you have guessed that this figure represents the spectrum of the planet Mars. Note that the figure includes the Doppler effect (as item 6), which we discuss next.

MA™ **The Doppler Effect Tutorial, Lessons 1, 2**

# 5.5 THE DOPPLER EFFECT

You're probably already amazed at the volume of information contained in light, but there is still more. In particular, we can learn about the motion of distant objects (relative to us) from changes in their spectra caused by the **Doppler effect.**

MATHEMATICAL INSIGHT 5.2

### Laws of Thermal Radiation

The two rules of thermal radiation have simple mathematical formulas. The *Stefan-Boltzmann law* (Law 1) is expressed as

$$\text{emitted power (per square meter of surface)} = \sigma T^4$$

where $\sigma$ (Greek letter *sigma*) is a constant with a measured value of $\sigma = 5.7 \times 10^{-8}\ \text{watt}/(\text{m}^2 \times \text{K}^4)$ and $T$ is on the Kelvin scale (K).

*Wien's law* (Law 2) is expressed approximately as

$$\lambda_{\text{max}} \approx \frac{2{,}900{,}000}{T\ (\text{Kelvin scale})}\ \text{nm}$$

where $\lambda_{\text{max}}$ (read as "lambda-max") is the wavelength (in nanometers) of maximum intensity, which is the peak of a thermal radiation spectrum.

**EXAMPLE:** Consider a 15,000 K object that emits thermal radiation. How much power does it emit per square meter? What is its wavelength of peak intensity?

**SOLUTION:**

**Step 1 Understand:** We can calculate the total emitted power per square meter from the Stefan-Boltzmann law and the wavelength of maximum intensity from Wien's law.

**Step 2 Solve:** We plug the object's temperature ($T$ = 15,000 K) into the Stefan-Boltzmann law to find the emitted power per square meter:

$$\sigma T^4 = 5.7 \times 10^{-8} \frac{\text{watt}}{\text{m}^2 \times \text{K}^4} \times (15{,}000\ \text{K})^4$$
$$= 2.9 \times 10^9\ \text{watt}/\text{m}^2$$

We find the wavelength of maximum intensity with Wien's law:

$$\lambda_{\text{max}} \approx \frac{2{,}900{,}000}{15{,}000\ (\text{Kelvin scale})}\ \text{nm} \approx 190\ \text{nm}$$

**Step 3 Explain:** A 15,000 K object emits a total power of 2.9 billion watts per square meter. Its wavelength of maximum intensity is 190 nm, which is in the ultraviolet portion of the electromagnetic spectrum. By using these facts in reverse, we can learn about astronomical objects. For example, if an object has a thermal radiation spectrum that peaks at a wavelength of 190 nm, we know that its surface temperature is about 15,000 K. And because the temperature tells us how much power the object emits per square meter of surface, we can calculate its total size from its total radiated power.