

4 THE ORIGIN AND NATURE OF LIGHT

IN THIS CHAPTER YOU WILL DISCOVER

- the origins of electromagnetic radiation
- the structure of atoms
- that stars with different surface temperatures emit different intensities of electromagnetic radiation
- that astronomers can determine the chemical compositions of stars and interstellar clouds by studying the wavelengths of electromagnetic radiation that they absorb or emit
- how to tell whether an object in space is moving toward or away from Earth

Stellar Spectra

Starlight passing through a prism or diffraction grating spreads into its component colors. From such spectra we can learn an incredible amount about stars, including their masses, surface temperatures, diameters, chemical compositions, and motions toward or away from us. This image, called a spectrogram, shows the spectrum of the Sun sliced and stacked to fit on this page. In it you can see thousands of absorption lines. As you study astronomy, note the distinct differences between the intensities of the various colors emitted by different stars. (NOAO)

R | V | U | X | G

WHAT DO YOU THINK?

- 1 How hot is a “red hot” object compared to objects glowing other colors?
- 2 What color is the Sun?
- 3 Can carbon dating be used to determine the age of space debris found on Earth?



Seeing objects in space is one thing, understanding their physical properties—temperature, chemical composition, mass, size, and motion—is something else altogether. Because almost all our knowledge about space comes from electromagnetic radiation, we now delve deeper into the nature of light. By understanding how light and matter interact, astronomers glean an amazing amount of information about the heavenly bodies.

BLACKBODY RADIATION

We begin our study of the interaction between light and matter with a perfectly familiar property of matter—temperature.

4-1 An object’s peak color shifts to shorter wavelengths as it is heated

Imagine that you have taken hold of an iron rod in a completely darkened room. You cannot see the rod, but you can feel its warmth in your hands. The rod is emitting enough infrared radiation (heat) for you to feel, but too few visible photons for your eyes to see. Now imagine that there is also

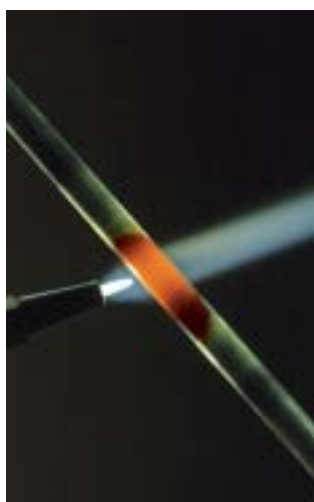
a propane torch in the room. You light it and heat the iron rod for several seconds, until it begins to glow (Figure 4-1a).

The rod’s first visible color is red—it glows “red hot.” Heated a little more, the rod appears orange and brighter (Figure 4-1b). Heated more, it appears yellow and brighter still (Figure 4-1c). After even more heating, the rod appears white-hot and brighter yet. If it doesn’t melt, further heating will make the rod appear blue and even brighter. This “thought experiment,” confirmed by the photographs in Figure 4-1, shows how the light any object emits changes with its temperature:

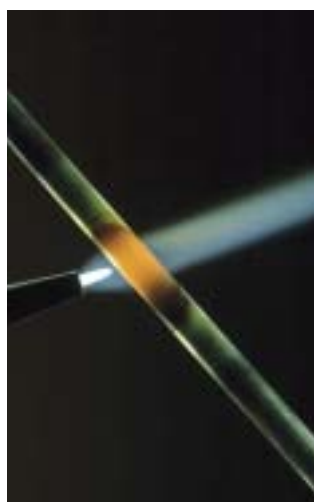
1. As an object heats up, it gets brighter because it emits more electromagnetic radiation.
2. The dominant color or wavelength of the emitted radiation changes with temperature. A cool object emits most of its energy at long wavelengths, such as infrared or red. A hotter object emits most of its energy at shorter wavelengths, such as blue, violet, or even ultraviolet.

Although the iron rod appears to emit a single color depending on its temperature, it actually gives off all wavelengths of electromagnetic radiation. So do stars, rocks, animals, and most other things in nature. However, a single wavelength of emitted radiation, denoted λ_{max} , dominates. When the object is cool, like a rock or animal, λ_{max} is a radio or infrared wavelength. When it is hot enough, λ_{max} is in the range of visible light, giving a hot object its characteristic color. Sufficiently hot stars emit λ_{max} in the ultraviolet part of the electromagnetic spectrum.

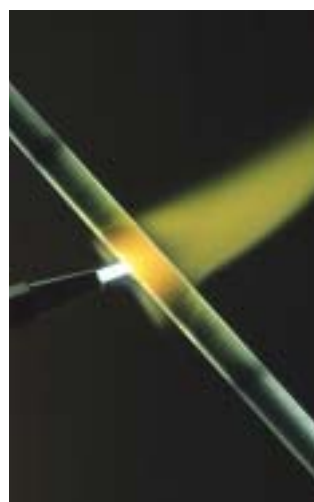
Iron bars and stars are good approximations to an important class of theoretical objects that astrophysicists call **blackbodies**. A blackbody absorbs all the electromagnetic radiation that strikes it. None of the incoming radiation is



a



b



c

R I V U X G

FIGURE 4-1 Heating a Bar of Iron This sequence of photographs shows the changing appearance of a piece of iron as it is heated. As the temperature increases, the amount of energy radiated by the bar increases. The apparent color of the bar also changes because, as the temperature goes up, the dominant wavelength of light emitted by the bar decreases. (© 1984 Richard Megna Fundamental Photographs)

reflected or scattered off its surface. The incoming radiation heats up the blackbody, which then reemits the energy it has absorbed, but not with the same intensities at each wavelength as it received.

ATM 4.3 *Intensity* measures how much energy a blackbody emits per second from any area on its surface. The intensities vary with the wavelength or color, and studies of blackbodies reveal that *the intensity of each wavelength emitted by a blackbody is determined solely by its temperature*. Profiles of the intensity of blackbody radiation at different wavelengths, called **blackbody curves**, are shown for three temperatures in Figure 4-2. (Temperatures throughout this book are expressed in Kelvins. If you are not familiar with the Kelvin temperature scale, review Toolbox A-1, “Temperature Scales,” at the back of the book.)

Unlike our iron rod, stars produce their own electromagnetic radiation rather than just absorbing and reradiating light from an outside source. Even so, stars behave like blackbodies, and the self-generated radiation they emit closely follows the idealized blackbody curves. We can, therefore, measure a real star’s surface temperature from the location of the peak of its intensity curve. To do that, we need to make our observations quantitative, that is, to express them mathematically.

4-2 The intensities of different emitted colors reveal a star’s temperature

ATM 4.5 **Wien’s law** is the mathematical relationship between the location of the color peak for each curve in Figure 4-2 and that blackbody’s temperature. In 1893 the German physicist Wilhelm Wien found that

The dominant wavelength of radiation emitted by a blackbody is inversely proportional to its temperature.

In other words, the hotter any object becomes, the shorter its λ_{max} , and vice versa.

Wien’s law proves very useful in computing the surface temperature of a star, because all we need to know is the dominant wavelength of the star’s electromagnetic radiation—we don’t need to know the star’s size, distance, or any other physical property.

A blackbody curve describes in detail the electromagnetic radiation a star emits. In 1879 the Austrian physicist Josef Stefan observed that

An object emits energy at a rate proportional to the fourth power of its temperature in Kelvins.

In other words, if you double the temperature of an object (for example, from 500 to 1000 K), the energy emitted from each square meter of the object’s surface each second in-

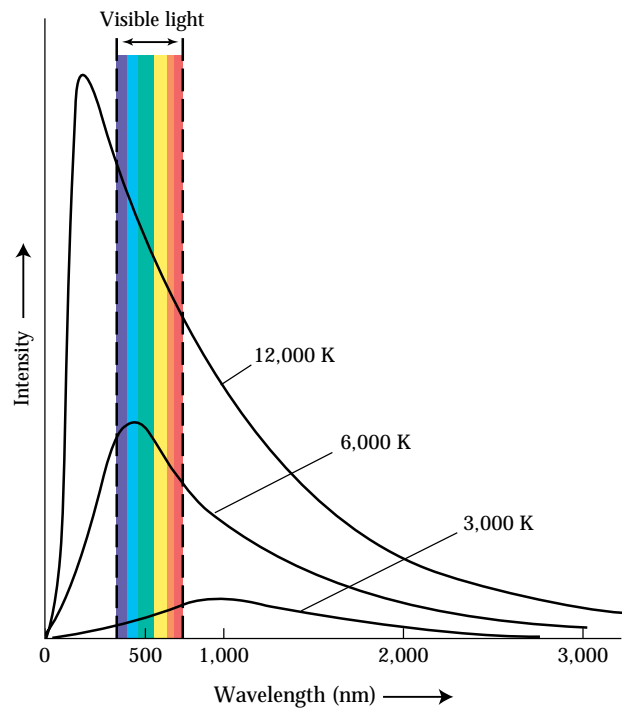


FIGURE 4-2 Blackbody Curves Three representative blackbody curves are shown here. Each curve shows the intensity of radiation at every wavelength emitted by a blackbody at a particular temperature. Note that hotter objects emit more of every wavelength and that the peak of emitted radiation changes with temperature.

creases by a factor of 2^4 , or 16 times. If you triple the temperature (for example, from 500 to 1500 K), the rate at which energy is emitted increases by a factor of 3^4 , or 81 times. Stefan’s experimental results were put on firm theoretical ground by Ludwig Boltzmann in 1885. The intensity-temperature relationship for blackbodies is named the **Stefan-Boltzmann law** in their honor.

The Stefan-Boltzmann law and Wien’s law are powerful tools for understanding the stars. If you would like to learn how to use these two radiation laws to make predictions, see An Astronomer’s Toolbox 4-1. These laws describe only two basic properties of *blackbody radiation*, namely, the relative rate at which it is emitted and the location of its peak. Because blackbodies give off all wavelengths of electromagnetic radiation, the blackbody curves, such as those already shown in Figure 4-2, give a more complete picture. These curves show the intensities of electromagnetic radiation emitted at different wavelengths by a blackbody with a given temperature. As we will see in the next section, the *differences* between an ideal blackbody curve and the curves seen from actual stars reveal stellar chemistries and motions of stars toward or away from us, among other things.

AN ASTRONOMER'S TOOLBOX 4-1

The Radiation Laws

Wien's law can be stated as a simple equation. If λ_{max} , the wavelength of maximum intensity, is measured in meters, then

$$\lambda_{\text{max}} = \frac{2.9 \times 10^{-3}}{T}$$

where T is the temperature of the blackbody measured in Kelvins.

Example: The Sun's maximum intensity is at a wavelength of about 500 nm, or 5×10^{-7} m. From Wien's law, we can calculate the Sun's surface temperature as

$$T = \frac{2.9 \times 10^{-3}}{5 \times 10^{-7}} = 5800 \text{ K}$$

The total energy emitted from each square meter of an object's surface each second is called the **energy flux**, F . In this context, flux means "rate of flow." Using this concept, we can write the Stefan-Boltzmann law as

$$F = \sigma T^4$$

where T is the temperature of the object in Kelvins and σ (lowercase Greek sigma) is a number called the Stefan-Boltzmann constant:

$$\sigma = 5.67 \times 10^{-8} \text{ J}/(\text{m}^2 \cdot \text{K}^4 \cdot \text{s})$$

where J is the energy unit joules, m is the unit meters, K is degrees Kelvin, and s is the unit seconds.

Because F is the energy emitted per second from each square meter of an object, multiplying F by the surface area, $4\pi r^2$, where r is the object's radius, yields the

total energy emitted by the object each second. Denoted L , and called **luminosity**, we have

$$L = F \cdot 4\pi r^2 = \sigma T^4 4\pi r^2$$

Luminosity is the total energy emitted per second by an entire object.

Compare! The Stefan-Boltzmann law has several applications. Suppose you observe two stars of equal size, one with a surface temperature of 10,000 K and the other with the same surface temperature as the Sun (5800 K). You can use the Stefan-Boltzmann law to determine how much brighter the hotter star is:

$$\frac{F_{\text{hotter}}}{F_{\text{colder}}} = \frac{\sigma \times 10,000^4}{\sigma \times 5800^4} \approx 8.8$$

So, the hotter star emits 8.8 times as much energy from each square meter of its surface than does the cooler star. Because the stars are the same size, the hotter one is therefore 8.8 times brighter than the cooler one.

Try these questions: The color yellow is centered around 550 nm. If the Sun were actually yellow, what would be its surface temperature? To find out why it is not yellow, see Guided Discovery: The Color of the Sun on page 100. What is the peak wavelength in nanometers given off by a blackbody at room temperature of 300 K? Referring to Figure 3-4, what part of the electromagnetic spectrum is that in?

(Answers appear at the end of the book.)

2

Figure 4-3 shows how the intensity of sunlight varies with wavelength. The blackbody curve for a body with a temperature of 5800 K is also plotted in Figure 4-3. Note how the observed intensity curve for the Sun closely follows the ideal blackbody curve at most wavelengths. Because the observed intensity curves for most stars and the idealized blackbody curves are so closely correlated, the laws of blackbody radiation can be applied to starlight. The peak of the intensity curve for the Sun is at a wavelength of about 500 nm, which is in the blue-green part of the visible spectrum.

By the end of the nineteenth century, physicists realized that they had reached an impasse in understanding electromagnetic radiation. While intensities and peak wavelengths were understood, all attempts to explain the characteris-

tic shapes of blackbody curves had failed. A breakthrough finally came in 1900, when the German physicist Max Planck derived a mathematical formula for the blackbody curves. He assumed that electromagnetic radiation is emitted in separate packets of energy. Light, a wave, behaves like a beam of particles. It was a remarkable result verified in 1905 by Albert Einstein, who called these particles of light **photons**.

We can now explain our thought experiment with the iron bar in terms of the properties of photons. Recall from Chapter 3 that the energy carried by a photon of light is inversely proportional to its wavelength. In other words, long-wavelength photons, such as radio waves, carry little energy. Short-wavelength photons, like X rays and gamma

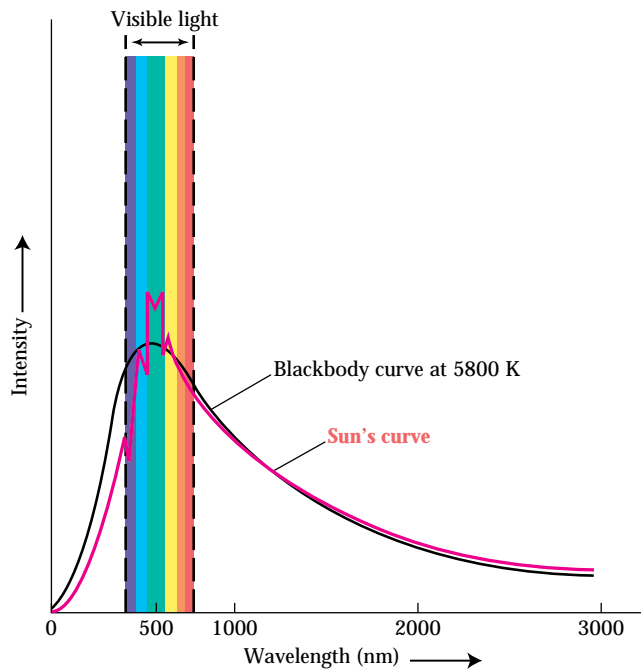


FIGURE 4-3 The Sun as a Blackbody This graph compares the intensity of sunlight (in red) over a wide range of wavelengths with the intensity of radiation from a blackbody at a temperature of 5800 K. The Earth's atmosphere scatters shorter-wavelength photons more than it scatters longer-wavelength ones. As seen from Earth, this scattering lowers the intensity of sunlight and shifts the blackbody peak to the right from where it actually occurs. To avoid these effects, the measurements of the Sun's intensity in this figure were made in space. The Sun mimics a blackbody remarkably well.

rays, carry much more energy. The relationship between the energy of a photon and its wavelength is called **Planck's law**. (You may want to refer back to An Astronomer's Toolbox 3-1.)

Together, Planck's law and Wien's law relate the temperature of an object to the energy of the photons it emits (Table 4-1). A cool object emits primarily long-wavelength photons that carry little energy, while a hot object gives off mostly short-wavelength photons that carry much more energy. In later chapters, we will find these ideas invaluable for understanding how stars of various temperatures interact with gas and dust in space.

DISCOVERING SPECTRA

In 1814, the German optician Joseph von Fraunhofer repeated Newton's classic experiment of shining a beam of sunlight through a prism (recall Figure 3-1). But Fraunhofer magnified the resulting rainbow-colored spectrum. He discovered that the solar spectrum contains hundreds of fine dark lines, which became known as **spectral lines**. Fraunhofer counted more than 600 such lines, and today physicists have detected more than 30,000 of them. Thousands of spectral lines are visible in the photograph of the Sun's spectrum shown in the figure opening this chapter.

By the mid-1800s, chemists discovered that they could produce spectral lines in the laboratory. Around 1857, the German chemist Robert Bunsen invented a special gas burner that produces a clean, colorless flame. Certain chemicals are easy to identify by the distinctive colors emitted when bits of the chemical are sprinkled into the flame of a Bunsen burner.

TABLE 4-1 Some Properties of Electromagnetic Radiation

	Wavelength (nm)	Photon energy (eV)*	Blackbody temperature (K)
Radio	$>10^7$	$<10^{-4}$	<0.03
Microwave**	10^7 to 4×10^5	10^{-4} to 3×10^{-3}	0.03 to 30
Infrared	4×10^5 to 7×10^2	3×10^{-3} to 2	30 to 4100
Visible	7×10^2 to 4×10^2	2 to 3	4100 to 7300
Ultraviolet	4×10^2 to 10^1	3 to 10^3	7300 to 3×10^6
X ray	10^1 to 10^{-2}	10^3 to 10^5	3×10^6 to 3×10^8
Gamma ray	$<10^{-2}$	$>10^5$	$>3 \times 10^8$

Note: $>$ means greater than, $<$ means less than.

*1 eV = 1.6×10^{-19} J.

**Microwaves, listed here separately, are often classified as radio waves or infrared radiation.

GUIDED DISCOVERY The Color of the Sun

Different people perceive the Sun to have different colors. To many it appears white, to others yellow. Still others, who notice it at sunset, believe it to be orange or even red. But we have seen that the Sun actually gives off all colors. Moreover, the peak in the Sun's spectrum falls between blue and green. Why doesn't the Sun appear turquoise? Several factors affect our perception of its color.

Before reaching our eyes, visible sunlight passes through the Earth's atmosphere. Certain wavelengths are absorbed and reemitted by the molecules in the air, a process called *scattering*. Violet light is scattered most strongly, followed in decreasing order by blue, green, yellow, orange, and red. That means that more violet, blue, and green photons are scattered by the Earth's atmosphere than are yellow photons. The intense scattering of violet, blue, and green has the effect of shifting the peak of the Sun's intensity entering our eyes from blue-green toward yellow. (The sky is blue because of the strong scattering of blue light—it isn't violet because the Sun emits many fewer violet photons than blue ones.)

The perception of a yellow Sun is further enhanced by our eyes themselves. Our eyes do not see all colors equally well. Rather, the light-sensitive cones in our eyes

each respond to one of three ranges of colors, which are centered on red, yellow, and blue wavelengths. None of the cones is especially sensitive to blue-green photons. By adding together the color intensities detected by the three types of cones, our brains *recreate* color. After combining all the light it can, the eye is most sensitive to the yellow-green part of the spectrum. We see blue and orange less well, and violet and red most poorly. Although the eye sees yellow and green light about equally well, the dominant color from the Sun is yellow because the air scatters green light, and our eyes are relatively insensitive to blue-green. Therefore, a casual glance at the Sun leaves the impression of a yellow object.

A longer look is extremely dangerous. **Don't try it!** Hypothetically, such a glance would leave the impression of a white Sun. The Sun's light is so intense that it would saturate the color-sensitive cones in our eyes; our brains interpret such saturation of the cones as white.

At sunrise and sunset we see an orange or red Sun. This occurs because close to the horizon the Sun's violet, blue, green, and even yellow photons are strongly scattered by the thick layer of atmosphere through which they travel, leaving the Sun looking redder and redder as it sets.

Bunsen's colleague, Gustav Kirchhoff, suggested that light from the colored flames could best be studied by passing it through a prism (Figure 4-4). The chemists promptly discovered that the spectrum from a flame consists of a pattern of thin, bright spectral lines against a dark background (not a blackbody spectrum!). They next found that *each chemical element produces its own characteristic pattern of spectral lines*. Thus was born in 1859 the technique of **spectral analysis**, the identification of chemical substances by their spectral lines.

4-3 Each chemical element produces its own unique set of spectral lines

A chemical **element** is a fundamental substance because it cannot be broken down into more basic units and still retain its properties. By the mid-1800s, chemists had already identified such familiar elements as hydrogen, oxygen, carbon, iron, gold, and silver. Spectral analysis promptly led to the discovery of additional elements, many of which are quite rare.

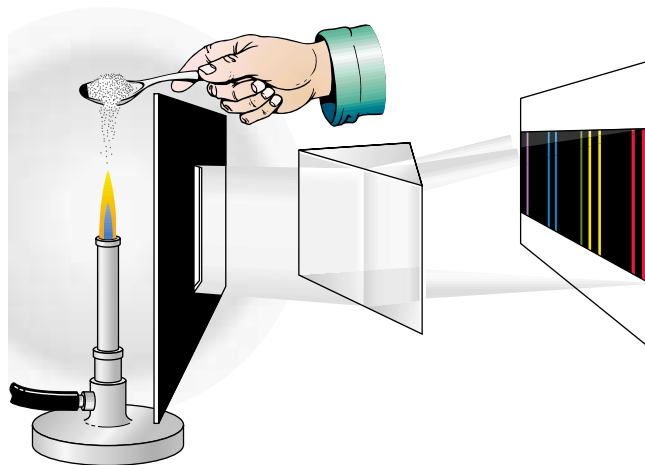


FIGURE 4-4 The Kirchhoff-Bunsen Experiment In the mid-1850s, Kirchhoff and Bunsen discovered that when a chemical substance is heated and vaporized, the resulting spectrum exhibits a series of bright spectral lines. In addition, they found that each chemical element produces its own characteristic pattern of spectral lines.

After Bunsen and Kirchhoff recorded the prominent spectral lines of all the known elements, they began to discover other spectral lines in mineral samples. In 1860, for example, they found a new line in the blue portion of the spectrum of mineral water. After chemically isolating the previously unknown element responsible for the line, they named it cesium (from the Latin *caesius*, meaning gray-blue). The next year a new spectral line in the red portion of the spectrum of a mineral sample led to the discovery of the element rubidium (from *rubidus*, for red).

During a solar eclipse in 1868, astronomers found a new spectral line in the light coming from the upper atmosphere of the Sun when the main body of the Sun was hidden by the Moon. This line was attributed to a new element, which was named helium (from the Greek *helios*, meaning Sun). Helium was not actually discovered on Earth until 1895, when it was identified in gases obtained from a uranium compound.

A list of the chemical elements is most conveniently displayed in the form of a **periodic table** (Figure 4-5). Each element has a unique atomic number (described in Section 4-5), and the elements are arranged in the periodic table by their atomic numbers. With a few exceptions, this sequence also corresponds to increasing average mass of the atoms of the elements. Thus, hydrogen (symbol

H) with atomic number 1 is the lightest element. Iron (Fe) has atomic number 26 and is a moderately heavy element.

All the elements in a single vertical column of the periodic table have similar chemical properties. For example, the elements in the far right column are all gases at room temperature and normal air pressure, and they rarely react chemically with other elements.

In addition to the 92 naturally occurring elements, Figure 4-5 lists the artificially produced elements. All the human-made elements are heavier than uranium (U), and all are highly radioactive, meaning that they spontaneously decay into lighter elements shortly after being created in the laboratory.

Insight into Science **Seek relationships** Finding how properties relate often provides invaluable insights into how new things work. The periodic table, for example, enables scientists to determine the properties of similar chemical elements. Keep an eye out for other such relationships throughout this book.

Because each chemical element produces its own unique pattern of spectral lines, scientists can determine the chemical composition of a remote astronomical object by identifying

Periodic Table of the Elements

1 H Hydrogen																	2 He Helium
3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
11 Na Sodium	12 Mg Magnesium											13 Al Aluminum	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
55 Cs Cesium	56 Ba Barium	57 La Lanthanum	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
87 Fr Francium	88 Ra Radium	89 Ac Actinium	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110	111	112		114		116		
58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium				
90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium				

FIGURE 4-5 Periodic Table of the Elements The periodic table is a convenient listing of the elements arranged according to their atomic numbers and chemical properties. Due to naming disagreements, elements 110–112, 114, and 116 do not yet have universally accepted names. Elements 113, 115, 117, and beyond have not yet been discovered.



R I V U X G

FIGURE 4-6 Iron in the Sun's Atmosphere The upper spectrum is a portion of the Sun's spectrum from 420 to 430 nm. Numerous dark spectral lines are visible. The lower spectrum is a corresponding portion of the spectrum of vaporized iron. Several

bright spectral lines can be seen against the black background. The fact that the iron lines coincide with some of the solar lines proves that there is some iron (albeit a very tiny amount) in the Sun's atmosphere. (Carnegie Observatories)

the lines in its spectrum. For example, Figure 4-6 shows a portion of the Sun's spectrum along with the spectrum of an iron sample taken here on Earth. No other chemical has iron's particular pattern of spectral lines. It is iron's own distinctive "fingerprint." Because the spectral lines of iron also appear in the Sun's spectrum, we can reasonably conclude that the Sun's atmosphere contains some vaporized iron.

Bunsen and Kirchhoff collaborated in designing and constructing the first **spectroscope**. This device consists of a prism and several lenses that magnify the spectrum so that it can be closely examined. After photography was invented, scientists preferred to produce a permanent photographic record of spectra. A device for photographing a spectrum is called a **spectrograph**, and this instrument is among the astronomer's most important tools.

In its basic form, a spectrograph consists of a slit, two lenses, and a prism arranged to focus the spectrum of an astronomical object onto a CCD (see Figures 3-27 or 3-28), as shown in Figure 4-7. This optical device typically mounts at the focal point of a telescope. The image of the object to be examined is focused on the slit. After the spectrum of a star or galaxy has been photographed, the CCD collects light from a known source, such as a gas composed of helium, neon, and argon, that is focused on the slit. This "comparison spectrum" is placed next to the spectrum of the object,

as in Figure 4-6. Because the wavelengths of the spectral lines in the comparison spectrum are already known from laboratory experiments, these lines can be used to identify and measure the wavelengths of the lines in the spectrum of the star or galaxy under study.

There are drawbacks to this type of spectrograph. A prism does not spread colors evenly: The blue and violet portions of the spectrum are spread out more than the red portion. In addition, because the blue and violet wavelengths must pass through more glass than the red wavelengths (see Figure 3-1), light is absorbed unevenly across the spectrum. Indeed, a glass prism is opaque to ultraviolet wavelengths.

A better device for breaking starlight into the colors of the rainbow is a **diffraction grating**, a piece of glass on which thousands of closely spaced lines are cut. Some of the finest diffraction gratings have as many as 10,000 lines per centimeter. The spacing of the lines must be very regular. Light rays reflected from different parts of the diffraction grating interfere with each other to produce a spectrum. Figure 4-8 shows the design of a modern diffraction grating spectrograph.

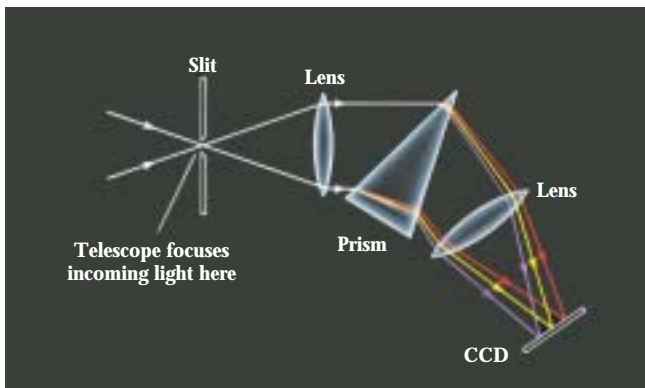


FIGURE 4-7 A Prism Spectrograph This optical device uses a prism to break up the light from an object into a spectrum. The lenses focus that spectrum onto a CCD.

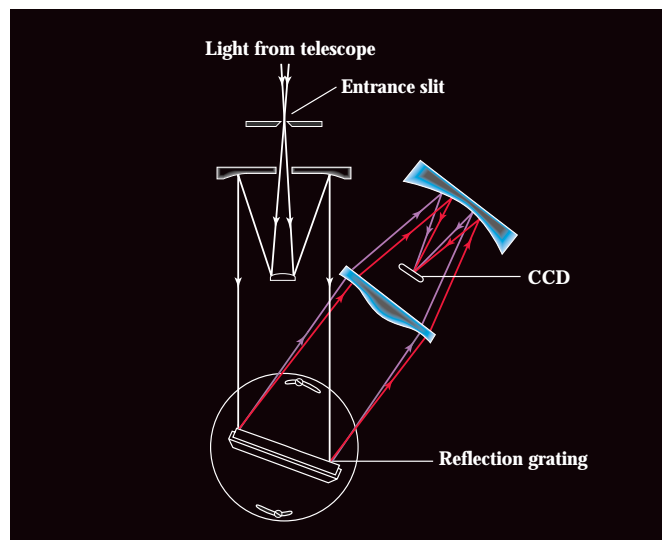


FIGURE 4-8 A Grating Spectrograph This optical device uses a grating to break up the light from an object into a spectrum. An arrangement of lenses and mirrors focuses that spectrum onto a CCD.

The spectral data are converted by computer to a graph that plots light intensity against wavelength. Dark lines in the rainbow-colored spectrum appear as depressions or valleys on the graph, while bright lines in the spectrum appear as peaks. For example, Figure 4-9 shows both a picture and a plot of a spectrum for hydrogen in which five dark spectral lines appear.

4-4 The brightnesses of spectral lines depend on conditions in the spectrum's source



During his pioneering experiments with spectra, Kirchhoff sometimes saw dark spectral lines, called *absorption lines*, among the colors of the rainbow. In other experiments, he saw bright spectral lines, called *emission lines*, against an otherwise dark background (see Figure 4-4). By the early 1860s, Kirchhoff had discovered the conditions under which these different types of spectra are observed. His description is summarized today as **Kirchhoff's laws**:

Law 1 A hot object or a hot, dense gas produces a **continuous spectrum** (also called a **continuum**)—a complete rainbow of colors without any spectral lines. This is a blackbody spectrum.

Law 2 A hot, rarefied gas produces an **emission line spectrum**—a series of bright spectral lines against a dark background.

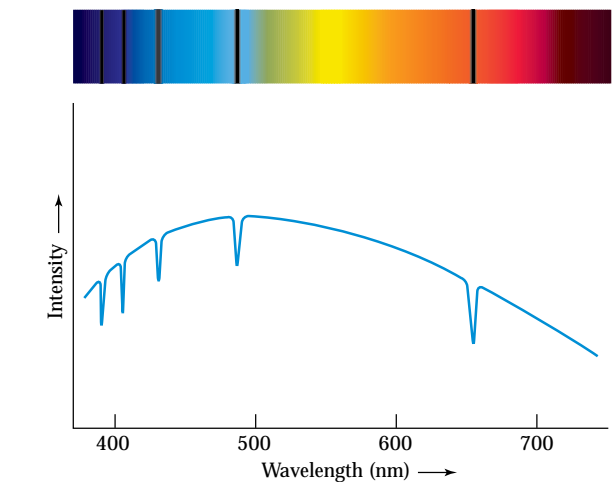


FIGURE 4-9 Spectrum of Hydrogen Gas (a) When a CCD is placed at the focus of a spectrograph, a rainbow-colored spectrum is recorded. (b) The spectrum is converted by computer into a graph of intensity versus wavelength. Note that the dark spectral lines appear as dips in the intensity-versus-wavelength curve.

Law 3 A cool gas in front of a continuous source of light produces an **absorption line spectrum**—a series of dark spectral lines among the colors of the rainbow.

To test your understanding of the above, you may want to try Interactive Exercise 4-1 before examining Figure 4-10, which summarizes how absorption and emission lines are formed.

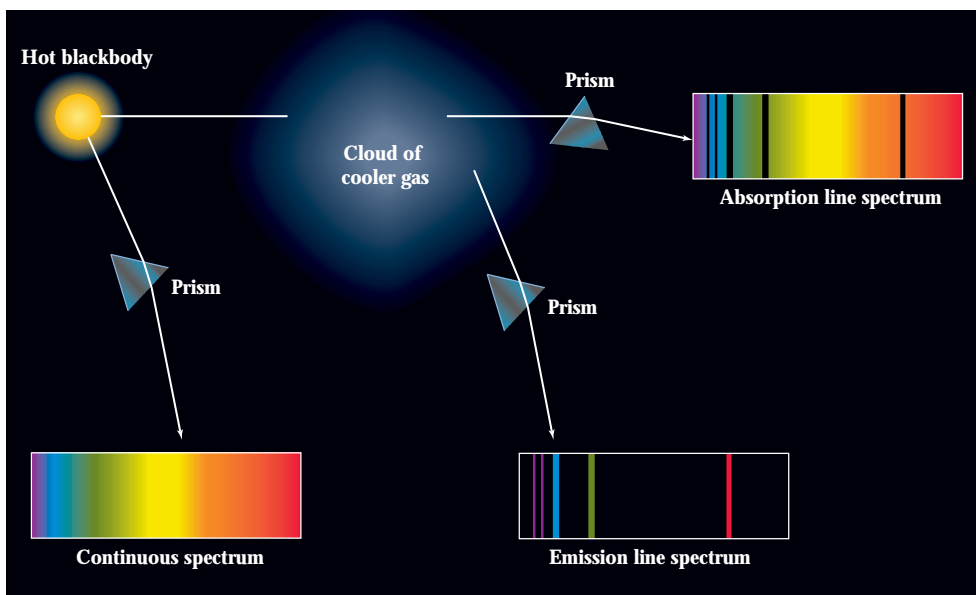


FIGURE 4-10 Continuous, Absorption Line, and Emission Line Spectra

This schematic diagram summarizes how different types of spectra are produced. A hot, glowing object emits a continuous spectrum. If this source of light is viewed through a cool gas, dark absorption lines appear in the resulting spectrum. When the same gas is viewed against a cold, dark background, its spectrum consists of just bright emission lines. The animation link shows the details of the activity drawn here.

Absorption lines are seen if the background is hotter than the gas. Emission lines are seen if the background is cooler. Note that the bright lines in the emission spectrum of a particular gas occur at exactly the same wavelengths as the dark lines in the absorption spectrum of that gas.

Consider, for example, the spectrum of the Sun. We know that the Sun's surface emits a continuous, blackbody spectrum, but here on Earth many absorption lines are seen in it (see the figure that opens this chapter). Kirchhoff's third law explains why. There must be a cooler gas between the surface of the Sun and the Earth. In fact, there are two: the Sun's lower atmosphere and the Earth's entire atmosphere. With these observations in hand, we turn now to exploring why the different spectra occur.

ATOMS AND SPECTRA

An atom is the smallest particle of a chemical element that still has the properties of that element. At the time of Kirchhoff's discoveries, scientists knew that all matter is composed of atoms, but they did not know their structures. Scientists saw that atoms of a gas somehow extract light of specific wavelengths from white light that passes through the gas, leaving dark absorption lines, and they concluded that the atoms then radiate light of precisely the same wavelengths—the bright emission lines (see Figure 4-6). But traditional theories of electromagnetism could not explain this phenomenon. The answer came early in the twentieth century with the development of quantum mechanics and nuclear physics.

4-5 An atom consists of a small, dense nucleus surrounded by electrons

The first important clue about the internal structure of atoms came in 1910 from Ernest Rutherford, a chemist and physicist from New Zealand. Rutherford and his colleagues at the University of Manchester in England were investigating the recently discovered phenomenon of radioactivity. Over time, a radioactive element naturally and spontaneously transforms into another element by emitting particles. Certain radioactive elements, such as uranium and radium, were known to emit such particles with considerable speed. It seemed plausible that a beam of these high-speed particles could penetrate a thin sheet of gold. Rutherford and his associates found that almost all the particles passed through the gold sheet with little or no deflection. To the surprise of the experimenters, however, an occasional particle bounced right back. It must have struck something very dense indeed.

Rutherford was quick to realize the implications of his experiment. He correctly reasoned that most of the mass of an atom is concentrated in a compact lump that occupies only a small part of the atom's volume. Most of the radioac-

tive particles pass freely through the nearly empty space that makes up most of the atom, but a few particles happen to strike the dense mass at the center and rebound.

Rutherford proposed a new model for the structure of an atom. According to this model, a massive, positively charged **nucleus** at the center of the atom is orbited by tiny, negatively charged *electrons* (Figure 4-11). Rutherford concluded that at least 99.95% of the mass of an atom is concentrated in the nucleus, whose diameter is only about one ten-thousandth the diameter of the atom.

Further research revealed that the nucleus of an atom contains two types of particles: *protons* and *neutrons*. A proton has almost the same mass as a neutron, and each proton or neutron is about 2000 times more massive than an electron. A proton has a positive electric charge and a neutron has no charge. Because like charges repel each other, protons do not naturally stay bound together; they try to move as far away from each other as possible. Neutrons help keep the protons bound together. Conversely, opposite charges attract, and this attraction keeps electrons in orbit around the nucleus, held there by the positively charged protons.

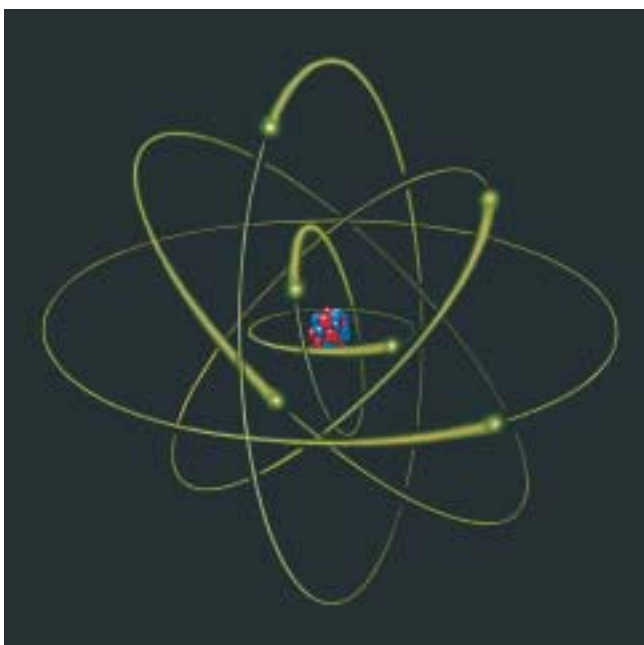


FIGURE 4-11 Rutherford's Model of the Atom Electrons orbit the atom's nucleus, which contains most of the atom's mass. The nucleus contains two types of particles: protons and neutrons. Because the nucleus and the electrons are so small compared to the distance between them, atoms are mostly empty space. The electrons, traveling at a few percent of the speed of light, make the nuclei appear to be surrounded by clouds.

AN ASTRONOMER'S TOOLBOX 4-2

Radioactivity and the Ages of Objects



Radioactive Age-Dating

The isotopes of many elements are radioactive, meaning that the elements spontaneously transform into other elements. Each radioactive isotope has a distinctive *half-life*, the time it takes half of the initial concentration of the isotope to transform into another element. After two half-lives, a radioactive isotope is reduced to $\frac{1}{2} \times \frac{1}{2}$ or $\frac{1}{4}$ of its initial concentration (see the accompanying figure). Among the most important radioactive elements for determining the age of objects in astronomy is the isotope of uranium with 146 neutrons, ^{238}U . The half-life of ^{238}U decaying into lead is 4.5 billion years.

To determine the time since an object, such as a piece of space debris discovered on Earth, solidified, scientists estimate how much lead it had when it formed. Then they measure the amounts of uranium and lead it contains now. Subtracting the amount of original lead, they use the amount of uranium and lead with the graph to determine how long ago the object formed.

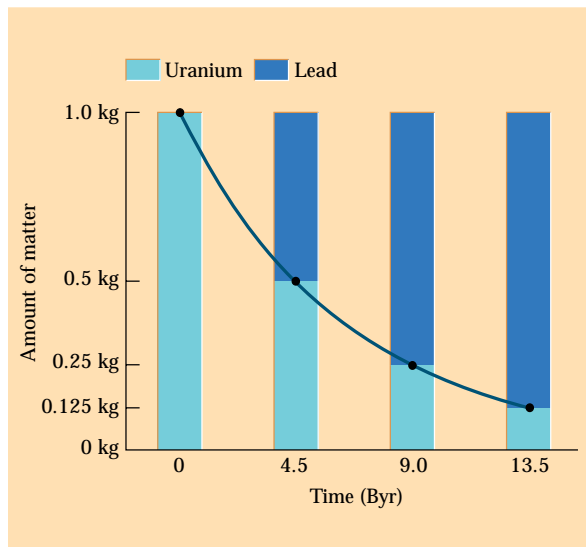
Example: Suppose a piece of space debris discovered on Earth was determined to have equal amounts of lead and uranium. How long ago did this debris form? Assuming that it originally had no lead, we see from the chart that a 1-to-1 mix of lead to uranium occurs 4.5 billion years after the object formed. This is one half-life of uranium.

3 Compare! This process works with any radioactive isotope. However, some isotopes have such short half-lives that they are not useful in astronomy. For example, the well-known carbon dating used to determine the ages of ancient artifacts on Earth is of little use in astronomy because ^{14}C has a half-life of only 5730 years. Because

so much of it has decayed away by then, carbon dating is useful for fewer than 100,000 years, usually a period over which little of astronomical importance occurs.

Try these questions: What fraction of a kilogram of radioactive material remains after 3 half-lives have passed? Using the figure below, estimate how much Uranium will remain after $6\frac{3}{4}$ billion years. Approximately how many half-lives of ^{14}C pass in 100,000 years?

(Answers appear at the end of the book.)



The Transformation of Uranium into Lead This figure shows the rate that 1 kg of uranium decays into lead, as described in the text. It contains $\frac{1}{8}$ kg of uranium after 13.5 billion years.

The number of protons in an atom's nucleus determines what element that atom is. Each element is assigned an **atomic number** that equals the number of protons it contains. Of the naturally occurring elements, a hydrogen nucleus always has 1 proton, a helium nucleus always has 2, and so forth, up to uranium with 92 protons in its nucleus (see Figure 4-5).

In contrast, the number of neutrons in the nuclei of different atoms of the same element may vary. For example, oxygen, the eighth element in the periodic table, with an atomic number of 8, always has eight protons, but it may

have eight, nine, or ten neutrons. These three slightly different kinds of oxygen are called **isotopes**. The isotope of oxygen with eight neutrons is by far the most abundant variety.

Many elements have isotopes that are radioactive. For example, carbon with six neutrons, ^{12}C , is stable, while carbon with eight neutrons, ^{14}C , is unstable. ^{14}C decays into nitrogen with seven neutrons, ^{14}N . To learn how radioactive decay is used to determine the ages of different objects, see An Astronomer's Toolbox 4-2.

Normally, the number of electrons orbiting an atom is equal to the number of protons in the nucleus, thus making



FIGURE 4-12 Balmer Lines in the Spectrum of a Star This portion of the spectrum of a star called HD 193182 shows more than two dozen Balmer lines. HD are the initials of Henry Draper. Draper and succeeding astronomers cataloged this and many other stars. The

series converges at 364.56 nm, just to the left of H_{40} . This star's spectrum also contains the first 12 Balmer lines (H_{α} through H_{12}), but they are not visible in this particular spectrogram. (Carnegie Observatories)

the atom electrically neutral. Astronomers denote neutral atoms by writing the atomic symbol followed by the roman numeral I. For example, neutral hydrogen is written as H I and neutral iron is Fe I.

When an atom contains a different number of electrons than protons, the atom is called an **ion**. The process of creating an ion is called **ionization**. Ions are denoted by the atomic symbol followed by a roman numeral that is one greater than the number of missing electrons. Positively ionized hydrogen (missing its one electron) is denoted H II, while positively ionized iron with seven electrons missing is denoted Fe VIII. It is worth noting that negative ions also exist, where nuclei have more electrons orbiting than they have protons. These are found, for example, in the Sun's outer layers and they help account for the Sun's blackbody curve (see Figure 4-3).

One way to create positive ions from neutral atoms is **photoionization**, in which photons of sufficiently high energy literally rip an electron completely out of orbit and into the space between atoms. Ionization occurs in stars, as seen by the fact that many of the spectral lines for stars correspond to ionized atoms.

Atoms can also share electrons and, by doing so, remain bound together. Such groups are called **molecules**. They are the essential building blocks of all complex structures, including life.

4-6 Spectral lines occur when an electron jumps from one energy level to another



The challenge of reconciling Rutherford's atomic model with the observations of spectral analysis was undertaken by the young Danish physicist Niels Bohr, who joined Rutherford's group at Manchester in 1911. Bohr began by trying to understand the structure of hydrogen, the lightest of the elements. The simplest hydrogen atom consists of a single electron and a single proton. Hydrogen has a visible spectrum consisting of a pattern of lines that begins at 656.28 nm and ends at 364.56 nm. The longest-

wavelength spectral line is called H_{α} , the second H_{β} , the third H_{γ} , and so forth, ending with the shortest-wavelength line, H_{∞} , at 364.56 nm. (The first few lines of the series are identified by Greek-letter subscripts; the remainder are identified by numerical subscripts.) The closer you get to 364.56 nm, the more spectral lines you see.

This spectral pattern had been described mathematically in 1885 by Johann Jakob Balmer, a German schoolteacher. By trial and error, Balmer discovered a formula for calculating the wavelengths of the hydrogen lines. Because of his discovery, the spectral lines of hydrogen at visible wavelengths are called **Balmer lines**, and the entire pattern from H_{α} to H_{∞} is called the **Balmer series**. The spectrum of the star shown in Figure 4-12 exhibits more than two dozen Balmer lines, from H_{13} through H_{40} .

Bohr's goal was to mathematically derive Balmer's formula from basic laws of physics. He began by assuming that the electron in a hydrogen atom moves around the nucleus only in certain specific orbits. As shown in Figure 4-13, it is customary to label these orbits $n = 1$, $n = 2$, $n = 3$, and so on. They are called the **Bohr orbits**.

Bohr argued that for an electron to jump, or to make a **transition**, from one orbit to another, the hydrogen atom must gain or lose a specific amount of energy. An electron jump from an inner orbit to an outer orbit requires energy; an electron jump from an outer orbit to an inner one releases energy. The energy gained or released by the atom when the electron changes orbits is the difference in energy between these two orbits.

According to Planck and Einstein, the packet of energy gained or released is a photon, whose energy is inversely proportional to its wavelength. Using these ideas, Bohr mathematically derived the formula that Balmer had discovered by trial and error. Furthermore, Bohr's discovery elucidated the meaning of the Balmer series: All the Balmer lines are produced by electron transitions between the second Bohr orbit ($n = 2$) and higher orbits ($n = 3, 4, 5$, and so forth). As an example, Figure 4-14 shows the electron transition that gives rise to the H_{α} spectral line, which has a wavelength of 656.28 nm.

In addition to giving the wavelengths of the Balmer series, Bohr's formula correctly predicts the wavelengths of other series of spectral lines that occur at nonvisible wavelengths. For example, electron transitions between the lowest Bohr orbit ($n = 1$) and all higher orbits also produce spectral lines. These transitions create the *Lyman series*, which is entirely in the ultraviolet wavelengths. At infrared wavelengths is the *Paschen series*, which arises out of transitions to and from the third Bohr orbit ($n = 3$). Additional series exist at still longer wavelengths.

Bohr's ideas also help explain Kirchhoff's laws. Each spectral line corresponds to one specific transition between the orbits of the electrons of a particular element. An absorption line is created when an electron jumps from an inner orbit to an outer orbit, extracting the required photon from an outside source of energy, such as the continuous spectrum of a hot, glowing object. An emission line is produced when an electron transitions to a lower orbit and emits a photon.

Today's view of the atom owes much to the Bohr model, but it is enhanced in certain ways. The modern theory of atoms is called **quantum mechanics**, a branch of physics dealing with photons and subatomic particles that was first developed during the 1920s. As a result of this work, physicists have discarded the concept that electrons are solid particles with planetlike orbits about the nucleus. Instead, electrons are now known to have both wave and particle properties; they are said to occupy certain allowed **energy levels** in the atom.

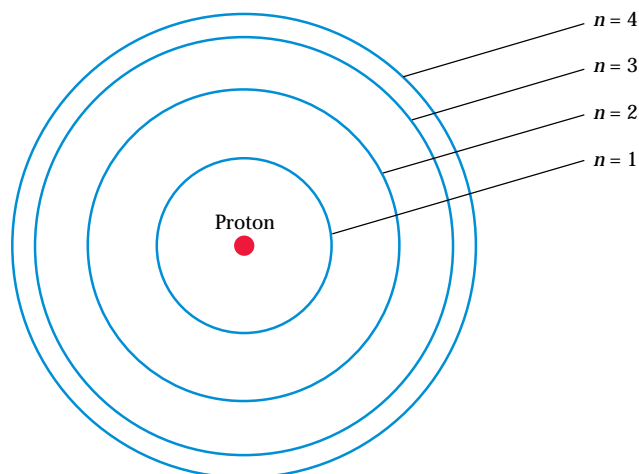


FIGURE 4-13 Bohr Model of the Hydrogen Atom According to Bohr's model of the atom, an electron circles the nucleus only in allowed orbits $n = 1, 2, 3$, and so on. The first four Bohr orbits are shown here.

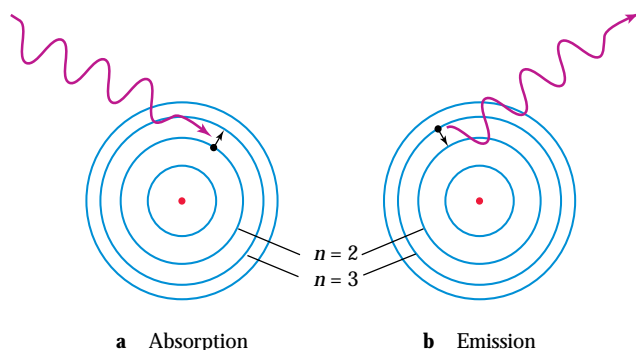


FIGURE 4-14 The Absorption and Emission of an H_{α} Photon This schematic diagram, drawn according to the Bohr model of the atom, shows what happens when a hydrogen atom absorbs or emits an H_{α} photon, which has a wavelength of 656.28 nm. (a) The photon is absorbed by the atom as the electron jumps from orbit $n = 2$ up to orbit $n = 3$. (b) The photon is emitted by the atom as the electron falls from orbit $n = 3$ down to orbit $n = 2$.

An extremely useful way of displaying the structure of an atom is an energy level diagram, such as that shown in Figure 4-15 for hydrogen. The lowest energy level, called the **ground state**, corresponds to the $n = 1$ Bohr orbit. An electron can jump from the ground state up to the $n = 2$ level only if the atom absorbs a photon of wavelength 121.6 nm. The energy of a photon is often expressed in electron volts (abbreviated eV). The 121.6 nm photon has an energy of 10.19 eV, so the energy level $n = 2$ is shown on the diagram as having an energy 10.19 eV above the energy of the ground state, which is usually assigned a value of 0 eV. Similarly, the $n = 3$ level is 12.07 eV above the ground state and so forth, up to the $n = \infty$ level at 13.6 eV. These states with more energy than the ground state are collectively called *excited states*.

If any electron in a hydrogen atom absorbs a photon with an energy greater than 13.6 eV, that electron will be knocked completely out of the atom. This is the process of photoionization mentioned above.

4-7 Spectral lines shift due to the relative motion between the source and the observer



Christian Doppler, a professor of mathematics in Prague, pointed out in 1842 that wavelength is affected by motion. As shown for the observer on

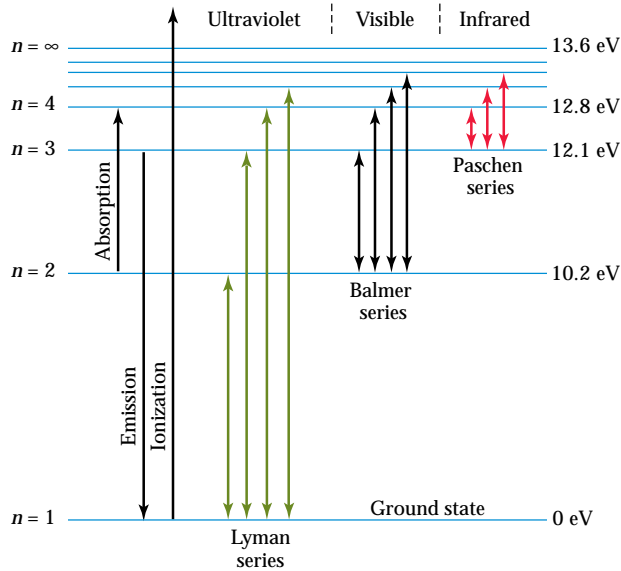


FIGURE 4-15 Energy Level Diagram of Hydrogen The structure of the hydrogen atom is conveniently displayed in a diagram showing the energy levels above the ground state. A variety of electron jumps, or transitions, are shown, including those that produce some of the most prominent lines in the hydrogen spectrum. For another perspective on the same information, you can try the linked Interactive Exercise.

the left in Figure 4-16, the wavelengths of electromagnetic radiation from an approaching source are compressed. The circles represent waves emitted from consecutive wave peaks as the source moves along. Because each successive wave is emitted from a position slightly closer to you, you see a shorter wavelength than you would if the source were stationary. All the spectral lines in the spectrum of an approaching source, such as a star, are therefore shifted toward the short-wavelength (blue) end of the spectrum, regardless of its distance. This phenomenon is called a **blueshift**.

Conversely, electromagnetic waves from a receding source are stretched out. You see a longer wavelength than you would if the source were stationary. All the spectral lines in the spectrum of a receding source, regardless of its distance, are shifted toward the longer-wavelength (red) end of the spectrum, producing a **redshift**. In either case, the effect of relative motion on wavelength is called the **Doppler shift**. This effect varies directly with approaching or receding speed: An object approaching twice as fast as another has its colors (or wavelengths) blueshifted twice as much when the approach speeds are small compared to the speed of light. An Astronomer’s Toolbox 4-3 gives a numerical example.

The speed determined from the Doppler shift is called the **radial velocity**, because the motion is along our line of sight, or along the “radius” drawn from Earth to the star. Of course, the star may well have motion perpendicular to our line of sight, across the celestial sphere. This **proper motion** does not affect the perceived wavelength and cannot be determined by Doppler shift. Proper motion is determined by measuring a star’s motion relative to background stars (Figure 4-17). The star with the greatest proper motion as seen from Earth is Barnard’s star (Figure 4-18).

Proper motions are so small that they can be measured only for relatively nearby stars in our Galaxy. However, the radial velocity of virtually every object in space can be determined regardless of their distances from us. For example, Doppler shift measurements of the spectra of hot gases on the Sun’s surface reveal that they rise and fall. Doppler shift measurements of stars in double star systems give crucial data about the speeds of the stars orbiting around each other. Doppler measurements of the spectra of distant galaxies enable us to determine the rate at which the entire universe is expanding. In later chapters, we will refer to the Doppler shift whenever we need to convert an observed wavelength shift into a speed.

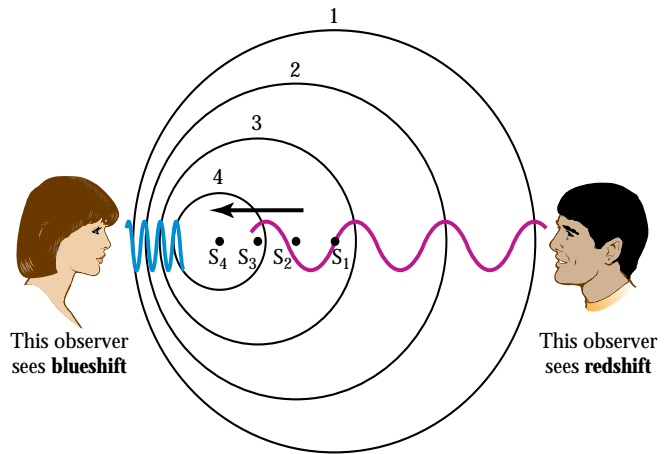


FIGURE 4-16 The Doppler Effect Wavelength is affected by motion between the light source and an observer. A source of light is moving toward the left. The four circles (numbered 1 through 4) indicate the location of light waves that were emitted by the moving source when it was at points S_1 through S_4 , respectively. Note that the waves are compressed in front of the source but stretched out behind it. Consequently, wavelengths appear shortened (blueshifted) if the source is moving toward the observer and lengthened (redshifted) if the source is moving away from the observer. Motion perpendicular to an observer’s line of sight does not affect wavelength.

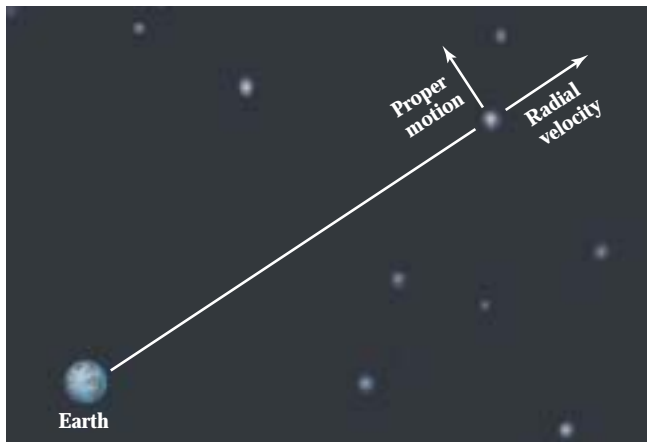


FIGURE 4-17 Radial and Proper Motions of a Star
The speed of a star toward or away from the Earth is its radial velocity. This motion creates a Doppler shift in the star's spectrum. The motion of the star across the sky, perpendicular to our line of sight, is called proper motion. Proper motion does not affect the star's spectrum.



R I V U X G

FIGURE 4-18 Barnard's Star This is a series of three superimposed photographs, taken over a four-year period. The line of three dots shows the proper motion of Barnard's star during that time. In addition to having the largest known proper motion (10.3" per year), Barnard's star is one of the closest stars to Earth. (John Sanford/Science Photo Library/Photo Researchers)

AN ASTRONOMER'S TOOLBOX 4-3

The Doppler Shift

Suppose that λ_0 is the wavelength of a spectral line from a stationary source. This is the wavelength that a reference book would list or a laboratory experiment would yield. If the source is moving, this particular spectral line is shifted to a different wavelength λ . The size of the wavelength shift is usually written as $\Delta\lambda$ (delta lambda), where $\Delta\lambda = \lambda - \lambda_0$. This $\Delta\lambda$ is the difference between the wavelength that you actually observe in the spectrum of a star or galaxy and the wavelength listed in reference books.

Doppler proved that the wavelength shift is governed by the simple equation

$$\Delta\lambda/\lambda_0 = v/c$$

where v is the speed of the source measured along the line of sight between the source and the observer. As usual, c is the speed of light (3×10^5 km/s).

Example: The spectral lines of hydrogen appear in the spectrum of the bright star Vega as shown in the figure. The prominent hydrogen line H_α has a normal wavelength of 656.285 nm, but in Vega's spectrum the line is located at 656.255 nm. The wavelength shift is -0.030 nm, so the star is approaching us with a speed of -14 km/s. The minus sign indicates that the star is moving toward us.

Try these questions: How fast would a star be moving toward us if all its wavelengths shifted to nine-tenths of their rest wavelengths? If a star is moving away from us at 100 km/s, what is the change, $\Delta\lambda/\lambda$? If a star is moving toward us at 0.1 c , what is the change, $\Delta\lambda/\lambda$, in its wavelengths?

(Answers appear at the end of the book.)



The Spectrum of Vega Several Balmer lines are seen in this photograph of the spectrum of Vega, the brightest star in the constellation Lyra (the Lyre). All the spectral lines are shifted equally and very slightly toward the blue side of the spectrum, indicating that Vega is approaching us. (NOAO)

Insight into Science Remote science Astronomical objects are so remote and their activities are often so complex that astrophysicists must use a tremendous amount of physics to interpret observations. For example, a single spectrum can contain information about stars, extrasolar planets, interstellar gas, the Earth's motion and atmosphere, and the performance of the observing telescope and the equipment attached to it. All of these factors must be understood theoretically and accounted for.

With the work of people like Planck, Einstein, Rutherford, Bohr, and Doppler, the interchange between astronomy and physics came full circle. Modern physics was born when Newton set out to understand the motions of the planets. Two and a half centuries later, physicists in their laboratories discovered the basic properties of electromagnetic radiation and the structures of atoms. As we will see in

WHAT DID YOU KNOW?

- 1 *How hot is a “red hot” object compared to objects glowing other colors?* Of all objects that glow visibly from heat stored or generated inside them, those that glow red are the coolest. Those that glow blue are the hottest.
- 2 *What color is the Sun?* The Sun emits all wavelengths of electromagnetic radiation. The colors it emits most intensely are in the blue-green part of the spectrum. Since the human eye is less sensitive to blue-green than

KEY WORDS

absorption line spectrum, 103
 atomic number, 105
 blackbody, 96
 blackbody curve, 97
 blueshift, 108
 continuous spectrum, 103
 diffraction grating, 102
 Doppler shift, 108
 element, 100
 emission line spectrum, 103
 energy flux, 98

energy level, 107
 ground state, 107
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 luminosity, 98
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proper motion, 108
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 radial velocity, 108
 redshift, 108
 spectral analysis, 100
 spectral lines, 99
 spectrograph, 102
 spectroscopy, 102
 Stefan-Boltzmann law, 97
 transition (of an electron), 106
 Wien's law, 97

KEY IDEAS

- By studying the wavelengths of electromagnetic radiation emitted and absorbed by an astronomical object, astronomers can learn the object's temperature, chemical composition, and movement through space.

later chapters, the fruits of their labors have important implications for astronomy even today.

4-8 Frontiers Yet to be Discovered

Spectra provide much of what we know about the cosmos. Spectra of objects in space continue to provide new insights into the chemical composition of the planets, moons, local space debris, stars, interstellar gas and dust, and galaxies, among other things. Furthermore, as we obtain more accurate Doppler measurements of objects in our Galaxy, in other nearby galaxies, and, indeed, of entire galaxies, we are better able to determine the rate at which the objects in the universe are moving relative to each other. The more we understand of this chemistry and motion, the more accurately we will be able to explain the evolution of the solar system, stars, galaxies, and the universe.



Further Reading on These Topics

- 3 *Can carbon dating be used to determine the age of space debris found on Earth?* No. Carbon dating is only reliable for objects that formed within the past 100,000 years. Space debris found on Earth formed more than 4.5 billion years ago.



Blackbody Radiation

- A blackbody is a hypothetical object that perfectly absorbs electromagnetic radiation at all wavelengths. Because a blackbody does not reflect any electromagnetic

radiation from outside sources, the radiation that it does emit depends only on its temperature. Stars closely approximate blackbodies.

- Wien's law states that the dominant wavelength of radiation emitted by a blackbody is inversely proportional to its temperature. The intensities of radiation emitted at various wavelengths by a blackbody at a given temperature are shown as a blackbody curve.
- The Stefan-Boltzmann law relates the temperature of a blackbody to the rate at which it radiates energy.

Discovering Spectra

- Spectroscopy—the study of electromagnetic spectra—provides important information about the chemical composition of remote astronomical objects.
- Kirchhoff's three laws of spectral analysis describe the conditions under which absorption lines, emission lines, and a continuous spectrum can be observed. Spectral lines serve as distinctive “fingerprints” for the chemical elements and chemical compounds comprising a light source.

Atoms and Spectra

- An atom consists of a small, dense nucleus (composed of protons and neutrons) surrounded by electrons that can occupy only certain energy levels.

- The spectral lines of a particular element correspond to the various electron transitions between allowed energy levels of that element. When an electron shifts from one energy level to another, a photon of the appropriate energy (and hence a specific wavelength) is absorbed or emitted by the atom.
- The spectrum of hydrogen at visible wavelengths consists of the Balmer series, which arises from electron transitions between the second energy level of the hydrogen atom and higher levels.
- When an atom loses or gains one or more electrons it is said to be charged, or ionized. One way for it to lose an electron is for the electron to absorb an energetic photon and thereby fly free from its atom.
- The spectral lines of an approaching light source are shifted toward shorter wavelengths (a blueshift); the spectral lines of a receding light source are shifted toward longer wavelengths (a redshift).
- The equation describing the Doppler effect states that the size of a wavelength shift is proportional to the radial velocity between the light source and the observer.

REVIEW QUESTIONS

- 1 What is a blackbody? What does it mean to say that a star behaves almost like a blackbody? If stars behave like blackbodies, why are they not black?
- 2 What is Wien's law? How could you use it to determine the temperature of a star's surface?
- 3 What is the Stefan-Boltzmann law? Why are astronomers interested in it?
- 4 Using Wien's law and the Stefan-Boltzmann law, explain the changes in color and intensity that are observed as the temperature of a hot, glowing object increases.

ADVANCED QUESTIONS

The answers to all computational problems, which are preceded by an asterisk (*), appear at the end of the book.

- *10 Approximately how many times around the world could a beam of light travel in 1 second?

- 5 Describe the experimental evidence that supports the Bohr model of the atom.
- 6 Explain how the spectrum of hydrogen is related to the structure of the hydrogen atom.
- 7 Why do different elements have different patterns of lines in their spectra?
- 8 What is the Doppler shift, and why is it important to astronomers?
- 9 Explain why the Doppler shift tells us only about the motion directly along the line of sight between a light source and an observer.

- *11 The bright star Regulus in the constellation of Leo (the Lion) has a surface temperature of 12,200 K. Approximately what is the dominant wavelength (λ_{max}) of the light it emits?

***12** The bright star Procyon in the constellation of Canis Minor (the Little Dog) emits the greatest intensity of radiation at a wavelength $\lambda_{\text{max}} = 445$ nm. Approximately what is the surface temperature of the star in Kelvins?

***13** The wavelength of H_{β} in the spectrum of the star Megrez in the Big Dipper is 486.112 nm. Laboratory measurements demonstrate that the normal wavelength of this spectral line is 486.133 nm. Is the star coming toward us or moving away from us? At what speed?

***14** In the spectrum of the bright star Rigel, H_{α} has a wavelength of 656.331 nm. Is the star coming toward us or moving away from us? How fast?

***15** Imagine driving down a street toward a traffic light. How fast would you have to go so that the red light would appear green?

DISCUSSION QUESTIONS

16 Compare the technique of identifying chemicals by their spectral line patterns with that of identifying people by their fingerprints.

17 Suppose you look up at the night sky and observe some of the brightest stars with your naked eye. Is there any way of telling which stars are hotter and which are cooler? Explain.

18 How can we exclude the Earth's atmosphere as the source of the iron absorption lines in Figure 4-6?

WHAT IF ...

***19** The Sun were twice its actual diameter, but still had the same surface temperature? At what wavelength would that new Sun emit its radiation most intensely? How many times brighter than the present Sun would the new Sun be? How might things on Earth be different under the bigger Sun?

20 All the nearby stars were observed to have redshifted spectra? What conclusions could we draw? What other measurement would be useful to have for each star and why?

21 No stars had any Doppler shift? What would that say about stellar motions relative to the solar system? Is that possible? *Clue:* Consider Newton's law of gravitation from Chapter 3.

22 The spectrum of the Sun (chapter-opening figure) had several absorption lines missing when taken from above the Earth's atmosphere compared to its spectrum taken at the Earth's surface?

WEB/CD-ROM QUESTIONS



***23 Measuring Stellar Temperatures** Access the Active Integrated Media Module "Blackbody Curves" in Chapter 4 of the *Discovering the Universe* Web site or CD-ROM. (a) Use the module to determine the range of temperatures over which a star's

peak wavelength is in the visible spectrum. (b) Determine if any of the following stars have a peak wavelength in the visible spectrum: Rigel, $T = 14,000$ K; Deneb, $T = 9500$ K; Arcturus, $T = 4500$ K; Vega, $T = 11,500$ K; Betelgeuse, $T = 3100$ K.

OBSERVING PROJECTS

24 As soon as weather conditions permit, observe a rainbow. Confirm that the colors are in the same order as shown in Figure 3-1a.

25 Obtain a glass prism or a diffraction grating, available from science museum stores, catalogs, or your college physics department. Look through the prism or grating at

various light sources, such as an ordinary incandescent lightbulb, a neon sign, and a mercury vapor street lamp. **Do not look at the Sun! Looking directly at the Sun causes blindness.** Do you have any trouble seeing spectra? What do you have to do to see a spectrum? Describe the differences in the spectra of the various light sources you observed.



26 Use *Starry Night Backyard*[™] to determine what planets, if any, are up at night now. Arrange to observe a planet and make an accurate drawing of its position among the background stars. If no planet is visible, you can perform these observations using the Moon. Observe the planet (or Moon) approximately one week later and again make a drawing. What type of motion have you plotted for the planet? Using the techniques in An Astronomer's Toolbox 1-1 or more accurate methods suggested by your instructor, determine the angle across the sky that the planet has moved in this time.



27 Use *Starry Night Backyard*[™] to examine three distant celestial objects. First view the entire sky (*Go:Atlas*). Then search for objects (i), (ii), (iii) listed below (*Edit:Find:object name*). Based on its appearance, state whether it is likely to have a continuum spectrum, an absorption line spectrum, or an emission line spectrum and explain your reasoning. (i) Trifid Nebula, (ii) M31, the Andromeda galaxy, (iii) the Moon. (*Hint*: Recall that moonlight is simply reflected sunlight.) Your teacher may assign additional objects to find.