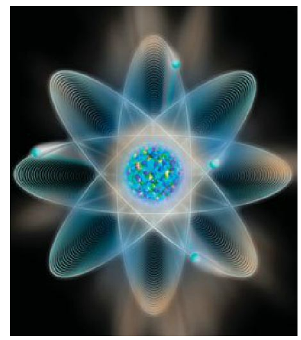
CHAPTER OUTLINE

THIS CHAPTER IN CONTEXT A good model is able to explain known facts and make useful predictions. In previous chapters, we've used a simple model of the atom as a collection of smaller particles—protons, neutrons, and electrons—to explain mass relationships between elements and to explain why isotopes exist. We have said that the number of protons in an atom's nucleus distinguishes it from atoms of other elements. For example, only sodium has 11 protons (and so, 11 electrons) in its atoms. Take away a proton and an electron, and you have neon, a colorless, odorless gas. Add a proton and an electron and you have magnesium, a lightweight metal that is much less reactive than sodium. In some way, then, the number of protons, neutrons, and electrons in an atom determines its properties.



This simple picture of the atom makes a nice corporate logo, but the idea of an atom with electrons orbiting a nucleus as planets orbit a sun was discarded nearly a century ago.

But how can viewing the atom as a core nucleus surrounded by electrons explain why one element is different from another? Why are some elements metals, while others are nonmetals? Why do metals form positive ions, while nonmetals form negative ones? Why do the properties of the elements repeat when they are arranged in order of increasing atomic number? And why do elements combine in certain ratios with other elements? Why, for example, is water's formula H_2O and not H_3O or HO? Our simple model of the atom cannot answer these questions.

In Chapter 1, we saw that experimental evidence indicates that atoms have a tiny, dense, positively charged core, the nucleus, surrounded by negatively charged electrons that fill the remaining volume of the atom. Electrons must move within this volume; otherwise, they would not be able to resist the electrostatic attraction of the nucleus, and the atom would collapse. However, a vibrating charge, such as an orbiting electron, creates ripples in the electric and magnetic fields around it. The creation of these ripples, or electromagnetic waves, should cause the electron to lose energy. As the electron's energy decreases, it should spiral inwards toward the nucleus. Calculations based on classical physics predict that after emitting a brief, intense burst of radiation, the electron should immediately crash into the nucleus (Figure 8.1). That simply doesn't happen. This **collapsing atom paradox** was only one sign that classical physics was severely flawed. An avalanche of experimental evidence convinced scientists that electrons do not behave at all like objects in the everyday world. An entirely new physics would be needed to explain how electrons behaved in atoms.

When experiment and theory do not agree, the theory must be modified or discarded. Classical physics could not resolve the collapsing atom paradox. It also could not correctly predict the heat capacity of cold solids, the periodicity of the elements, or crucial characteristics of the radiation emitted by hot objects or hot samples of the elements.

New experiments led to the startling realization that electrons were not simply tiny, negatively charged particles. While electrons do behave like particles in some experiments, they behave like waves in others. This **wave/particle duality** has no real analogy in everyday life. The physics of objects that exhibit wave/particle duality is called **quantum mechanics,** or **quantum theory.** Quantum theory is a cornerstone of modern chemistry.

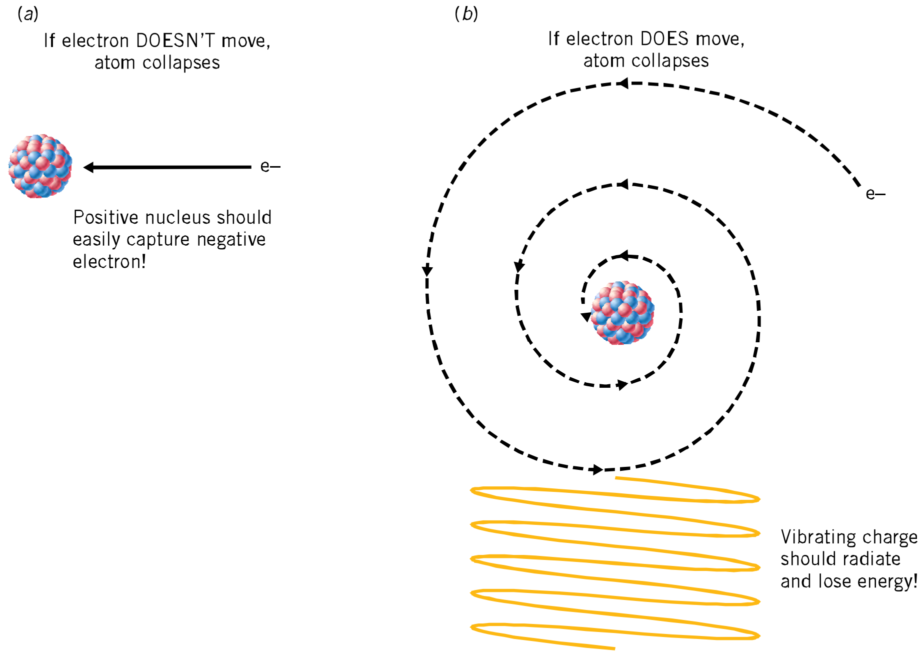


FIGURE 8.1 ***The collapsing atom paradox.*** (*a*) If the electron does not move, the positively charged nucleus will attract and capture it—and the atom will collapse. (*b*) If the electron moves around the nucleus, it should radiate energy as an electromagnetic wave. The electron spirals into the nucleus, and the atom collapses.

Quantum theory can be used to explain why atoms are stable, why things have the colors they do, why the periodic table has the shape it has, why chemical bonds form, and why some molecules form but others don't. Quantum theory played a key role in the rapid advances in electronics and communications technology that took place in the last century. In this century, it has profound importance in the development of new computing technologies and in nanotechnology, an emerging field concerned with the design and construction of molecular machines.

We'll discuss nanotechnology further in Chapter 13.

Both electromagnetic radiation and electrons exhibit wave/particle duality. To understand the experimental basis for the quantum theory, we have to begin our discussion with radiation.

8.1 Electromagnetic radiation can be described as a wave or as a stream of photons

Electromagnetic radiation can be described as a wave

You've learned that objects can have energy in only two ways, as kinetic energy and as potential energy. You also learned that energy can be transferred between things, and in Chapter 7 our principal focus was on the transfer of heat. Energy can also be transferred between things as light or radiation. This is a very important form of energy in chemistry. For example, many chemical systems emit visible light as they react (see Figure 8.2).

Electricity and magnetism are closely related to each other. A moving charge creates an electric current, which in turn creates a magnetic field around it. This is the fundamental idea behind electric motors. A moving magnetic field creates an electric field or current. This is the idea behind electrical generators and turbines.

Many experiments show that radiation carries energy through space by means of **waves.** Waves are an oscillation that moves outward from a disturbance (think of ripples moving away from a pebble dropped into a pond). In the case of radiation, the disturbance can be a vibrating electric charge. When the charge jiggles, it produces a pulse in the electric field around it. As the electric field pulses, it creates a pulse in the magnetic field. The magnetic field pulse gives rise to yet another electric field pulse further away from the disturbance. The process continues, with a pulse in one field giving rise to a pulse in the other, and the resulting train of pulses

(*a*) (*b*) (*c*)

FIGURE 8.2 ***Light is given off in a variety of chemical reactions.*** (*a*) Combustion. (*b*) Cyalume® light sticks. (*c*) A lightning bug.

in the electric and magnetic fields is called an **electromagnetic wave.** The wave ripples away from the source at almost unimaginable speeds.

The higher the vibrating charge bounces, the greater the height or **amplitude** of the peaks will be. Amplitude affects the intensity or brightness of the radiation. Figure 8.3 shows how the amplitude or intensity of the wave varies with time and with distance as the wave travels through space. In Figure 8.3*a*, we see two complete oscillations or *cycles* of the wave during a 1 second interval. The number of cycles per second is called the **frequency** of the electromagnetic radiation, and its symbol is ν (the Greek letter *nu*, pronounced "new").

Electromagnetic waves don't need a medium to travel through, as water and sound waves do. They can cross empty space. The speed of the electromagnetic wave in a vacuum is the same no matter how the radiation is created (about 3.00×10^8 m/s).

The concept of frequency extends beyond just electromagnetic radiation to other events that recur at regular intervals. For example, you go to school 5 days *per week,* or you pay your bills once *each month*. Each of these statements describes how frequently an event occurs, and they have in common the notion of *per unit of time,* or $\frac{1}{\text{time}}$. In the SI, the unit of time is the second (s), so frequency is given the unit "per second," which is $\frac{1}{\text{second}}$, or $(\text{second})^{-1}$. This unit is given the special name **hertz (Hz).**

The SI symbol for the second is s.

$$s^{-1} = \frac{1}{s}$$

$$1 \text{ Hz} = 1 \text{ s}^{-1}$$

As electromagnetic radiation moves away from its source, the positions of maximum and minimum amplitude (peaks and troughs) are regularly spaced. The peak-to-peak distance is called the radiation's **wavelength,** symbolized by λ (the Greek letter *lambda*). (See Figure 8.2*b*.) Because wavelength is a distance, it has distance units (e.g., meters).

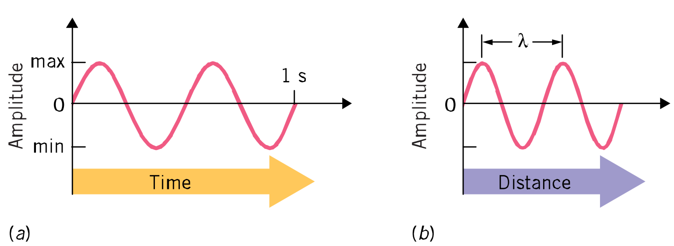


FIGURE 8.3 ***Two views of electromagnetic radiation.*** (*a*) The frequency, ν, of a light wave is the number of complete oscillations each second. Here two cycles span a 1 second time interval, so the frequency is two cycles per second, or 2 Hz. (*b*) Electromagnetic radiation frozen in time. This curve shows how the amplitude varies along the direction of travel. The distance between two peaks is the wavelength, λ, of the electromagnetic radiation.

For any wave, the product of its wavelength and its frequency equals the speed of the wave.

If we multiply the wavelength by frequency, the result is the speed of the wave. We can see this if we analyze the units.

$$\text{meters} \times \frac{1}{\text{second}} = \frac{\text{meters}}{\text{second}} = \text{speed}$$

In SI units

$$\text{m} \times \frac{1}{\text{s}} = \frac{\text{m}}{\text{s}} = \text{m s}^{-1}$$

The speed of light is one of our most carefully measured constants, because the meter is defined in terms of it. The precise value of the speed of light in a vacuum is 2.99792458×10^8 m/s, and a meter is defined as exactly the distance traveled by light in 1/299792458 of a second.

The speed of electromagnetic radiation in a vacuum is a constant and is commonly called the *speed of light*. Its value to three significant figures is 3.00×10^8 m/s (or m s^{-1}). This important physical constant is given the symbol c.

$$c = 3.00 \times 10^8 \text{ m s}^{-1}$$

From the preceding discussion we obtain a very important relationship that allows us to convert between λ and ν.

TOOLS

Wavelength-frequency relationship

$$\lambda \times \nu = c = 3.00 \times 10^8 \text{ m s}^{-1} \qquad (8.1)$$

EXAMPLE 8.1

Calculating Frequency from Wavelength

Mycobacterium tuberculosis, the organism that causes tuberculosis, can be completely destroyed by irradiation with ultraviolet light with a wavelength of 254 nm. What is the frequency of this radiation?

ANALYSIS: To convert between wavelength and frequency we use Equation 8.1. However, we must be careful about the units.

SOLUTION: To calculate the frequency, we solve Equation 8.1 for ν.

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

Next, we substitute for c (3.00×10^8 m s^{-1}) and for the wavelength. However, to cancel units correctly, we must have the wavelength in meters. Recall from Chapter 1 that nm means nanometer and the prefix nano implies the factor "$\times 10^{-9}$":

$$1 \text{ nm} = 10^{-9} \text{ m}$$

Therefore, 254 nm equals 254×10^{-9} m. Substituting gives

$$\nu = \frac{3.00 \times 10^8 \text{ m s}^{-1}}{254 \times 10^{-9} \text{ m}}$$
$$= 1.18 \times 10^{15} \text{ s}^{-1}$$
$$= 1.18 \times 10^{15} \text{ Hz}$$

Does the Answer Make Sense?

If the frequency is about 10^{15} Hz, the speed of light should be

$$(254 \times 10^{-9} \text{ m}) \times (10^{15} \text{ s}^{-1}) = 254 \times 10^{6} \text{ m s}^{-1} \approx 3 \times 10^8 \text{ m s}^{-1}$$

which is correct to one significant digit.

EXAMPLE 8.2

Calculating Wavelength from Frequency

Radio station WGBB on Long Island, New York, broadcasts its AM signal, a form of electromagnetic radiation, at a frequency of 1240 kHz. What is the wavelength of these radio waves expressed in meters?

ANALYSIS and SOLUTION: Once again we are relating wavelength and frequency, so we must use Equation 8.1. This time we solve Equation 8.1 for the wavelength.

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

Now we must cancel the unit s^{-1} in 3.00×10^8 m s^{-1}. The prefix "k" in kHz means kilo and stands for "$\times 10^3$" and Hz means s^{-1}.

$$1 \text{ kHz} = 10^3 \text{ Hz}$$

$$1 \text{ Hz} = 1 \text{ s}^{-1}$$

The frequency is therefore 1240×10^3 s^{-1}. Substituting gives

$$\lambda = \frac{3.00 \times 10^8 \text{ m s}^{-1}}{1240 \times 10^3 \text{ s}^{-1}}$$

$$= 242 \text{ m}$$

Does the Answer Make Sense?

If this wavelength is correct, we should be able to multiply it by the original frequency (in Hz) and get the speed of light, as we did in the previous example. We also should be able to divide the speed of light by the wavelength to get back the frequency. To two figures,

$$\frac{3.0 \times 10^8 \text{ m/s}}{2.4 \times 10^2 \text{ m}} = 1.2 \times 10^6 \text{ Hz} = 1200 \text{ kHz}$$

which is reasonably close to the original frequency of 1240 kHz.

PRACTICE EXERCISE 1: The most intense radiation emitted by the Earth has a wavelength of about 10.0 μm. What is the frequency of this radiation in hertz?

PRACTICE EXERCISE 2: An FM radio station in West Palm Beach, Florida, broadcasts electromagnetic radiation at a frequency of 104.3 MHz (megahertz). What is the wavelength of these radio waves, expressed in meters?

Electromagnetic waves are categorized by frequency

Electromagnetic radiation comes in a broad range of frequencies called the **electromagnetic spectrum,** illustrated in Figure 8.4. Some portions of the spectrum have popular names. For example, radio waves are electromagnetic radiations having very low frequencies (and therefore very long wavelengths). Microwaves, which also have low frequencies, are emitted by radar instruments such as those the police use to monitor the speeds of cars. In microwave ovens, similar radiation is used to heat water in foods, causing the food to cook quickly. Infrared radiation is emitted by hot objects and consists of the range of frequencies that can make molecules of most substances vibrate internally. You can't see infrared radiation, but you can feel how your body absorbs it by holding your hand near a hot radiator; the absorbed radiation makes your hand warm. Gamma rays are at the high-frequency end of the electromagnetic spectrum. They are produced by certain elements that are radioactive. X rays are very much like gamma rays, but they are usually made by special equipment. Both X rays and gamma rays penetrate living things easily.

Remember that there is an inverse relationship between wavelength and frequency. The lower the frequency, the longer the wavelength.

Most of the time, you are bombarded with electromagnetic radiations from all portions of the electromagnetic spectrum. Radio and TV signals pass through you; you feel infrared radiation when you sense the warmth of a radiator; X rays and gamma rays fall on you from space; and light from a lamp reflects into your eyes from the page you're reading. Of all these radiations, your eyes are able to sense only a very narrow band of wavelengths ranging from about 400 to 700 nm. This band is called the **visible spectrum** and consists of all the colors you can see, from red through orange, yellow, green, blue, and violet. White light is composed of all these colors in roughly equal amounts, and it can be separated into them by focusing a beam of white light through a prism, which spreads the various wavelengths

The speed of electromagnetic radiation was computed to be about 3×10^8 m/s. The fact that the same speed was determined experimentally for light supported the hypothesis that light was a form of electromagnetic radiation.

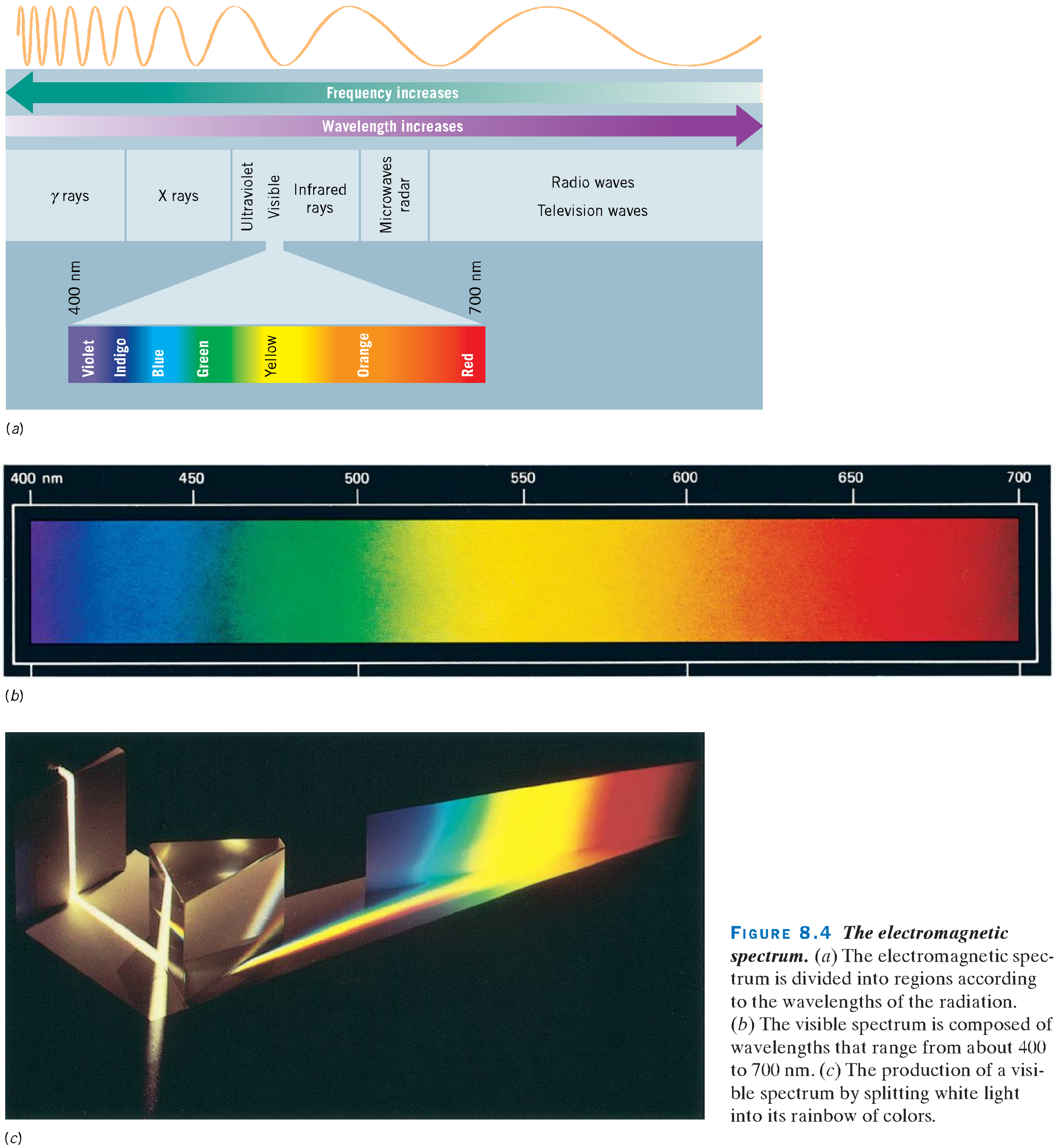


FIGURE 8.4 ***The electromagnetic spectrum.*** (*a*) The electromagnetic spectrum is divided into regions according to the wavelengths of the radiation. (*b*) The visible spectrum is composed of wavelengths that range from about 400 to 700 nm. (*c*) The production of a visible spectrum by splitting white light into its rainbow of colors.

apart. This is illustrated in Figure 8.4*b*. A photograph showing the production of a visible spectrum is given in Figure 8.4*c*.

The way substances absorb electromagnetic radiation often can help us characterize them. For example, each substance absorbs a uniquely different set of infrared frequencies. A plot of the wavelengths absorbed versus the intensities of absorption is called an *infrared absorption spectrum*. It can be used to identify a compound, because each infrared spectrum is as unique as a set of fingerprints. (See Figure 8.5.) Many substances absorb visible and ultraviolet radiations in unique ways, too, and they have visible and ultraviolet spectra (Figure 8.6).

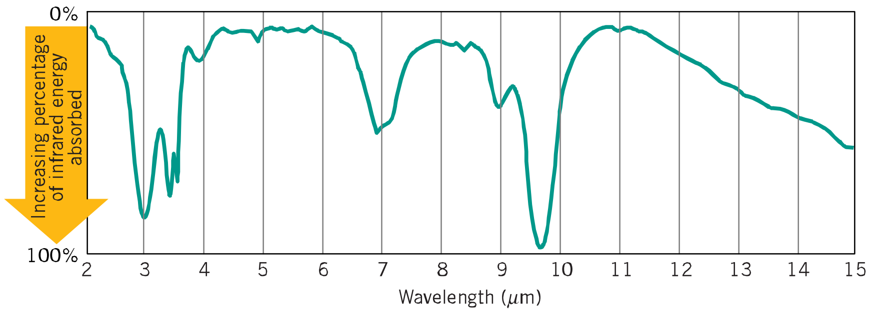


FIGURE 8.5 ***Infrared absorption spectrum of methyl alcohol (also called wood alcohol), the fuel in "canned heat" products such as Sterno.*** In an infrared spectrum, the usual practice is to show the amount of light absorbed increasing from top to bottom in the graph. Thus, there is a peak in the percentage of light absorbed at about 3 μm. (Spectrum courtesy Sadtler Research Laboratories, Inc., Philadelphia, Pa.)

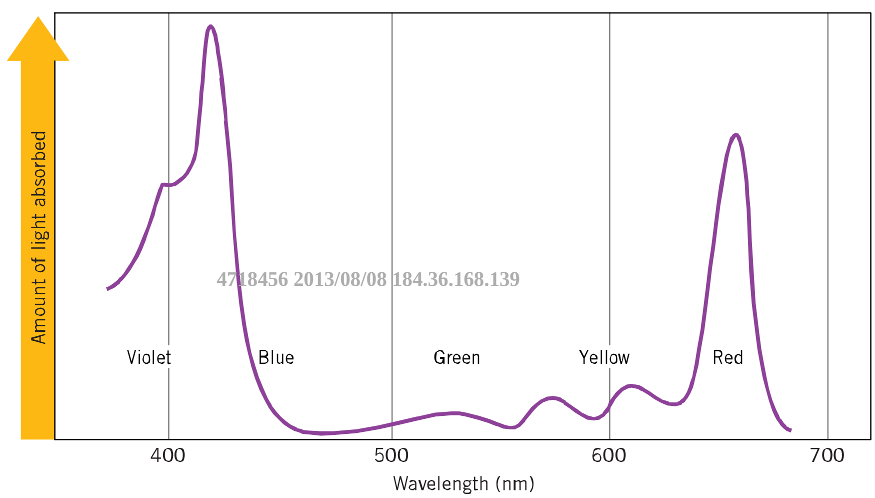
Electromagnetic radiation can be viewed as a stream of photons

When an electromagnetic wave passes an object, the oscillating electric and magnetic fields may interact with it, as ocean waves interact with a buoy in a harbor. A tiny charged particle placed in the path of the wave will be yanked back and forth by the changing electric and magnetic fields. For example, when a radio wave strikes an antenna, electrons within the antenna begin to bounce up and down, creating an alternating current that can be detected and decoded electronically. Because the wave exerts a force on the antenna's electrons and moves them through a distance, work is done. Thus, as energy is lost by the source of the wave (the radio transmitter), energy is gained by the electrons in the antenna.

A series of groundbreaking experiments showed that classical physics does not correctly describe energy transfer by electromagnetic radiation. In 1900, a German physicist named Max Planck (1858–1947) proposed that electromagnetic radiation can be viewed as a stream of tiny packets or **quanta** of energy that were later called **photons.** Each photon travels at the speed of light. Planck proposed, and Albert

The energy of one photon is called one **quantum** of energy.

(*a*)



(*b*)

FIGURE 8.6 ***Absorption of light by chlorophyll.*** (*a*) Chlorophyll is the green pigment plants use to harvest solar energy for photosynthesis. (*b*) In this visible absorption spectrum of chlorophyll, the percentage of light absorbed increases from bottom to top. Thus, there is a peak in the light absorbed at about 420 nm and another at about 660 nm. This means the pigment strongly absorbs blue-violet and red light. The green color we *see* is the light that's *not* absorbed. It's composed of the wavelengths of visible light that are reflected. (Our eyes are most sensitive to green, so we don't notice the yellow components of the reflected light.)

FACETS OF CHEMISTRY 8.1

Photoelectricity and Its Applications

One of the earliest clues to the relationship between the frequency of light and its energy was the discovery of the photoelectric effect. In the latter part of the nineteenth century, it was found that certain metals acquired a positive charge when they were illuminated by light. Apparently, light is capable of kicking electrons out of the surface of the metal.

When this phenomenon was studied in detail, it was discovered that electrons could only be made to leave a metal's surface if the frequency of the incident radiation was above some minimum value, which was named the *threshold frequency.* This threshold frequency differs for different metals, depending on how tightly the metal atom holds onto electrons. Above the threshold frequency, the kinetic energy of the emitted electron increases with increasing frequency of the light. Interestingly, however, its kinetic energy does not depend on the intensity of the light. In fact, if the frequency of the light is below the minimum frequency, no electrons are observed at all, no matter how bright the light is. To physicists of that time, this was very perplexing because they believed the energy of light was related to its brightness. The explanation of the phenomenon was finally given by Albert Einstein in the form of a very simple equation.

$$\text{KE} = h\nu - w$$

where KE is the kinetic energy of the electron that is emitted, $h\nu$ is the energy of the photon of frequency ν, and w is the minimum energy needed to eject the electron from the metal's surface. Stated another way, part of the energy of the photon is needed just to get the electron off the surface of the metal. This amount is w. Any energy left over ($h\nu - w$) appears as the electron's kinetic energy.

Besides its important theoretical implications, the photoelectric effect has many practical applications. For example, automatic "electric eye" door openers use this phenomenon by sensing the interruption of a light beam caused by the person wishing to use the door. The phenomenon is also responsible for photoconduction by certain substances that are used in light meters in cameras and other devices. The production of sound in motion pictures was first made possible by incorporating a strip along the edge of the film (called the *sound track*) that causes the light passing through it to fluctuate in intensity according to the frequency of the sound that's been recorded. A photocell converts this light to a varying electric current that is amplified and played through speakers in the theater. Even the sensitivity of photographic film to light is related to the release of photoelectrons within tiny grains of silver bromide that are suspended in a coating on the surface of the film.

Einstein (1879–1955) confirmed, that *the energy of a photon of electromagnetic radiation is proportional to the radiation's frequency,* not to its intensity or brightness as had been believed up to that time. (See Facets of Chemistry 8.1.)

TOOLS

Energy of a photon

The value of Planck's constant is 6.626×10^{-34} J s. It has units of energy (joules) multiplied by time (seconds).

$$\text{energy of a photon} = E = h\nu \qquad (8.2)$$

In this expression, h is a proportionality constant that we now call **Planck's constant.** Note that Equation 8.2 relates two representations of electromagnetic radiation. The right-hand side of the equation deals with a property of particles (energy per photon); the left-hand side deals with a property of waves (the frequency). Quantum theory unites the two representations, so we can use whichever representation of electromagnetic radiation is convenient for describing experimental results. For example, in describing the photoelectric effect (Facets of Chemistry 8.1), we represent radiation as a stream of particles. When describing how radiation can bend around small obstacles and fan out after passing through pinholes, we represent radiation as a wave phenomenon. Radiation is not a stream of particles, or a wave—it's something else entirely, something unlike anything we have ever encountered in everyday life. It is something that can behave like either particles or waves, depending on the experiment.

Planck's and Einstein's discovery was really quite surprising. If a particular event requiring energy, such as photosynthesis in green plants, is initiated by the absorption of light, it is the frequency of the light that is important, not its intensity or brightness. This makes sense when we view light as a stream of photons, since "high frequency" is associated with higher energy of the photons, while higher intensity is associated with greater numbers of photons. But it makes no sense at all when we view radiation as a wave phenomenon.

Brighter light delivers more photons; higher-frequency light delivers more energetic photons.

The idea that electromagnetic radiation can be represented as either a stream of photons or a wave is a cornerstone of the quantum theory. Physicists were able to use the concept of photons to understand many experimental results that classical physics simply had no explanation for. The success of the quantum theory in describing radiation paved the way for a second startling realization: that electrons, like radiation, could be represented as either waves or particles. We now turn our attention to the first experimental evidence that led to our modern quantum mechanical model of atomic structure: the existence of discrete lines in atomic spectra.

8.2 Atomic line spectra are experimental evidence that electrons in atoms have quantized energies

The visible spectrum described in Figure 8.4 is called a **continuous spectrum** because it contains a continuous unbroken distribution of light of *all* colors. It is formed when the light from the sun, or any other object that's been heated to a very high temperature (such as the filament in an electric light bulb), is split by a prism and displayed on a screen. A rainbow after a summer shower is a continuous spectrum that most people have seen. In this case, tiny water droplets in the air spread out the colors contained in sunlight.

A rather different kind of spectrum is observed if we examine the light that is given off when an *electric discharge,* or spark, passes through a gas such as hydrogen. The electric discharge is an electric current that *excites,* or energizes, the atoms of the gas. More specifically, the electric current excites and gives energy to the electrons in the atom. The atoms then emit the absorbed energy in the form of light as the electrons return to a lower energy state. When a narrow beam of this light is passed through a prism, as shown in Figure 8.7, we do *not* see a continuous spectrum. Instead, only a few colors are observed, displayed as a series of individual lines. This series of lines is called the element's **atomic spectrum** or **emission spectrum.** Figure 8.8 shows the visible portions of the atomic spectra of two common elements, sodium and hydrogen, and how they compare with a continuous spectrum. Notice that the spectra of these elements are quite different. In fact, each element has its own unique atomic spectrum that is as characteristic as a fingerprint.

Atoms of an element can also be excited by adding them to the flame of a Bunsen burner.

An emission spectrum is also called a ***line spectrum*** because the light corresponding to the individual emissions appears as lines on the screen.

A simple pattern of lines in the spectrum of hydrogen suggests a simple explanation for atomic spectra

The first success in explaining atomic spectra quantitatively came with the study of the spectrum of hydrogen. This is the simplest element, since its atoms have only one electron, and it produces the simplest spectrum with the fewest lines.

The atomic spectrum of hydrogen actually consists of several series of lines. One series is in the visible region of the electromagnetic spectrum and is shown in

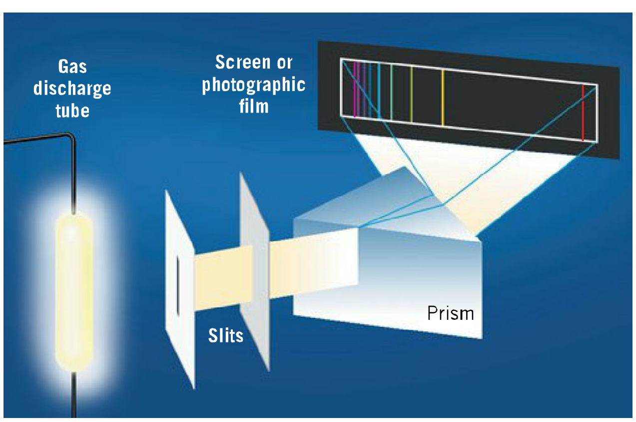


FIGURE 8.7 ***Production and observation of an atomic spectrum.*** Light emitted by excited atoms is formed into a narrow beam and passed through a prism, which divides the light into relatively few narrow beams with frequencies that are characteristic of the particular element that's emitting the light. When these beams fall on a screen, a series of lines is observed, which is why the spectrum is also called a *line spectrum.*

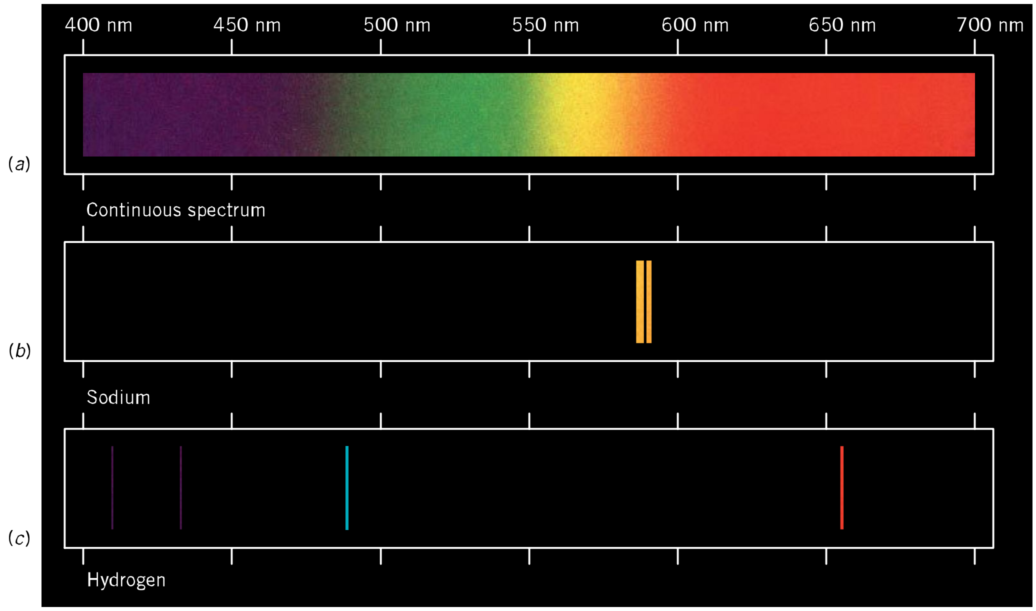
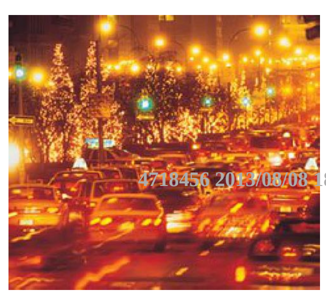


FIGURE 8.8 ***Continuous and atomic emission spectra.*** (*a*) The continuous visible spectrum produced by the sun or an incandescent lamp. (*b*) The atomic spectrum emission produced by sodium. The emission spectrum of sodium actually contains more than 90 lines in the visible region. The two brightest lines are shown here. All the others are less than 1% as bright as these. (*c*) The atomic spectrum (line spectrum) produced by hydrogen. There are only four lines in this visible spectrum. They vary in brightness by only a factor of 5, so they are all shown.

CHEMISTRY IN PRACTICE

At one time, most of the street lighting in towns and cities was provided by incandescent lamps, in which a tungsten filament is heated white-hot by an electric current. Unfortunately, much of the light from this kind of lamp is infrared radiation, which we cannot see. As a result, only a relatively small fraction of the electrical energy used to operate the lamp actually results in visible light. Modern streetlights consist of high-intensity sodium or mercury vapor lamps in which the light is produced by passing an electric discharge through the vapors of these metals. The electric current excites the atoms, which then emit their characteristic atomic spectra. In these lamps, most of the electrical energy is converted to light in the visible region of the spectrum, so they are much more energy-efficient (and cost-efficient) than incandescent lamps. Sodium emits intense light at a wavelength of 589 nm, which is yellow. The golden glow of streetlights in scenes like that shown in the photograph at the right is from this emission line of sodium, which is produced in high-pressure sodium vapor lamps. There is another, much more yellow light emitted by low-pressure sodium lamps that you may also have seen. Both kinds of sodium lamps are very efficient, and almost all communities now use this kind of lighting to save money on the costs of electricity. In some places, the somewhat less efficient mercury vapor lamps are still used. These lamps give a bluish-white light.

Fluorescent lamps also depend on an atomic emission spectrum as the primary source of light. In these lamps an electric discharge is passed through mercury vapor, which emits some of its light in the ultraviolet region of the spectrum. The inner wall of the fluorescent tube is coated with a phosphor that glows white when struck by the UV light. These lamps are also highly efficient and convert most of the electrical energy they consume into visible light.



Park Avenue in New York City is brightly lit by sodium vapor lamps in this photo taken during the Christmas season.