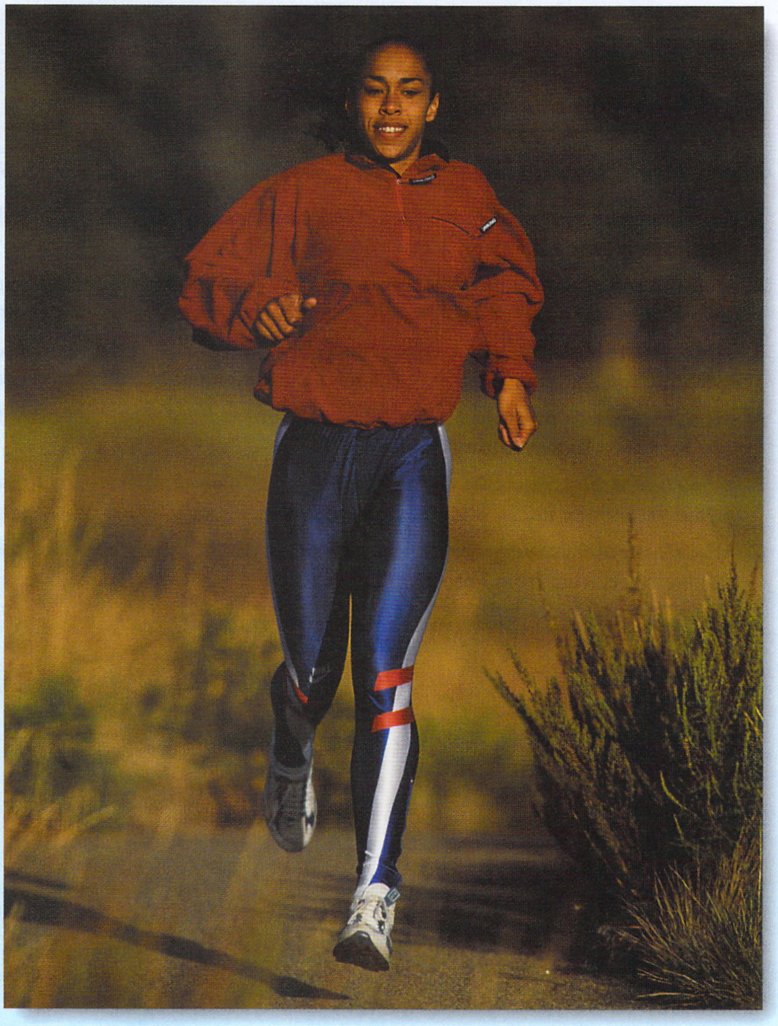
21

Atomic Structure and Interactions

KEY IDEA

The arrangement of protons and electrons in atoms determines their chemical properties.



PHYSICS AROUND US . . . A Deep Breath

Take a deep breath. As you feel the air moving into your lungs, pause for a moment to think about what it is you are taking into your body. You probably know that some of the air you are taking in is made from atoms of a gas called oxygen. These atoms are taken into your bloodstream and carried to your cells, where they participate in chemical reactions that provide you with the energy you need to live. These reactions are similar to what happens when a piece of wood burns in a fire—atoms of oxygen combine with atoms of carbon to produce a molecule called carbon dioxide, which is in your breath when you exhale. This carbon dioxide, in turn, is taken in by plants, whose chemical processes produce the oxygen that you breath in. Atoms of oxygen, in other words, move around through the biosphere. They appear in combination with different sorts of atoms at different times, but the oxygen atoms themselves don't change.

Another part of the air you breathe in (most of it, actually) is made from a different kind of atom, called nitrogen. Unlike oxygen, nitrogen doesn't take part in combustion, either inside or outside the body. Isn't it amazing that these two atoms, mingled so closely together, can have such different properties?

The next time you're out in the country breathing in that fresh air, think about the fact that the unchanging atoms you are drawing in are part of an endless cycle of combination and recombination—a cycle that makes life on our planet possible.

TURNING INWARD

By the end of the nineteenth century, the industrialized world was feeling pretty confident. The shimmering prospect of endless progress permeated public thought—after all, this was the generation that had seen the telegraph remove barriers to communication, electricity light up their cities, and trolley cars move people around with unprecedented speed. There seemed to be no limit to what human society could accomplish.

The accomplishments of physics seemed to fit this pattern. Newton's laws of motion explained the motion of the solar system and other material objects. Maxwell's equations codified and linked the phenomena of electricity and magnetism (and, incidentally, led to the availability of commercial electric power). The laws of thermodynamics explained the behavior of heat and gave engineers a way of improving the steam engines that had powered the Industrial Revolution (although, to be honest, many improvements came from plain old-fashioned tinkering).

And yet . . . and yet . . . there were still mysteries. They seemed small things—footnotes in the triumphal march—but they had something in common. As the century turned, scientists began to see more new and puzzling things in the behavior of the smallest bits of matter. Eventually, they turned their attention inward, toward the atom and its constituents. In the end, these new phenomena led to an entirely new way of looking at nature.

A Bewildering Variety of Atoms

In Chapter 9 we discussed John Dalton's atomic theory and how it slowly gained acceptance during the nineteenth century. The idea that each chemical element corresponded to a different kind of atom and that atoms could combine to form molecules explained all the observations and rules that scientists had made. However, there were still doubters, scientists who recognized that the concept of atoms explained many observations, but were still not certain that they existed in reality.

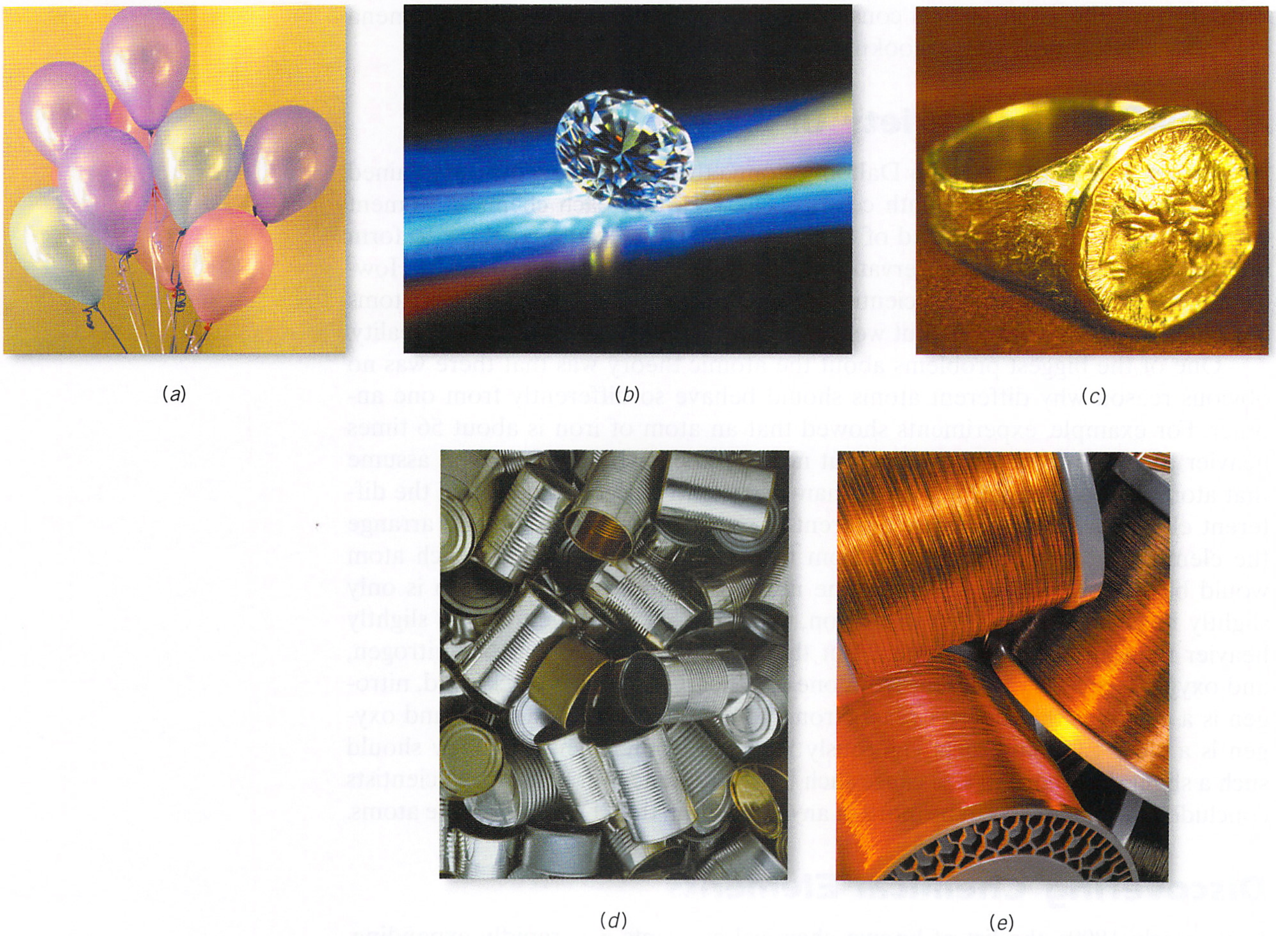
One of the biggest problems about the atomic theory was that there was no obvious reason why different atoms should behave so differently from one another. For example, experiments showed that an atom of iron is about 56 times heavier than an atom of hydrogen, but nobody knew why. Suppose you assume that atoms of iron are simply bigger than atoms of hydrogen and that all the different elements are just slightly different in size (and mass). You could arrange the elements according to weight, from lightest to heaviest, so that each atom would be slightly heavier than the one next to it. An atom of nitrogen is only slightly heavier than an atom of carbon, and an atom of oxygen is only slightly heavier than an atom of nitrogen. But these three elements—carbon, nitrogen, and oxygen—are very different from one another. Carbon is a black solid, nitrogen is a clear gas that doesn't react strongly with most other elements, and oxygen is a clear gas that reacts vigorously with most other elements. Why should such a slight difference in size cause such large differences in behavior? Scientists concluded that there had to be more of an explanation than Dalton's simple atoms.

Discovering Chemical Elements

In the early 1800s, the list of known chemical elements was rapidly expanding, but contained only a few dozen entries. Today, the periodic table lists more than 110 elements, of which 92 appear in nature and the rest have been produced artificially. Most of the materials we encounter in everyday life are not elements but compounds of two or more elements bound together. Table salt, plastics, stainless steel, paint, window glass, and soap are all made from a combination of elements.

Nevertheless, we do have experience with a few chemical elements in our everyday lives:

- **Helium:** A light gas that has many uses in addition to filling party balloons and blimps. In liquid form, helium is used to maintain superconductors at low temperatures (see Chapter 24).
- **Carbon:** Pencil lead, charcoal, and diamonds are all examples of pure carbon. The differences among these materials have to do with the way the atoms of carbon are linked together, as we discuss in Chapter 24.
- **Gold:** A soft, yellow, dense, and highly valued metal. For thousands of years the element gold has been coveted as a symbol of wealth. Today it coats critical electrical contacts in spacecraft and other sophisticated electronics.



(a) Helium in a party balloon; (b) carbon in diamonds; (c) gold in the tomb of King Tutankhamen; (d) aluminum in cans; and (e) copper in electrical wire.

- **Aluminum:** A lightweight metal used for many purposes, from overhead power lines to airplane parts and building construction. The dull white surface of the metal is actually a combination of aluminum and oxygen, but if you scratch the surface, the shiny material underneath is pure elemental aluminum.
- **Copper:** The reddish metal of pennies and pots. Copper wire provides a relatively inexpensive and efficient conductor of electricity.

Although we know of more than 90 different elements in nature, many natural systems are constructed from just a few. Six elements—oxygen, silicon, magnesium, iron, aluminum, and calcium—account for almost 99% of Earth's solid mass. Most of the atoms in your body are hydrogen, carbon, oxygen, or nitrogen, with smaller but important roles played by phosphorus and sulfur. And most stars are formed almost entirely from the lightest element, hydrogen. These differences in the behavior of the elements suggested that atoms might have a complex internal structure that made them different from one another.

THE INTERNAL STRUCTURE OF THE ATOM

Dalton's idea of the atom as a single indivisible entity was not destined to last. In 1897, English physicist Joseph John Thomson (1856–1940) unambiguously identified a particle called the **electron,** which has a negative electric charge and is much smaller and lighter than even the smallest atom known. Because there was no place from which a particle such as the electron could come, other than inside the atom, Thomson's discovery provided incontrovertible evidence for what some physicists had suspected for a long time: Atoms are not the fundamental building blocks of matter, but they are made up of particles that are smaller and more fundamental still. Table 21-1 summarizes some of the important terms related to atoms.

The Atomic Nucleus

The most important discovery about the structure of the atom was made by New Zealand–born physicist Ernest Rutherford (1871–1937) and his coworkers in Manchester, England, in 1911. The basic idea of the experiment is sketched in

TABLE 21-1 Important Terms Related to Atoms

Atom	The smallest particle that retains its chemical identity
Electron	A subatomic particle with negative charge and small mass
Nucleus	The small, massive central part of an atom
Proton	A positively charged nuclear particle
Neutron	An electrically neutral nuclear particle
Ion	An electrically charged atom
Element	A chemical substance made up of only one kind of atom
Molecule	Any collection of two or more atoms bound together

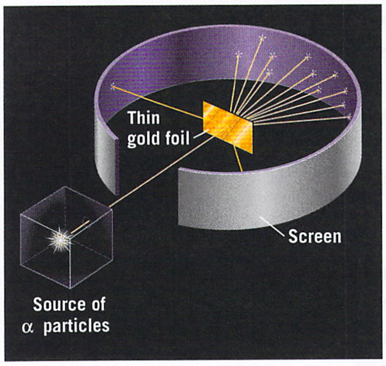


FIGURE 21-1. In Rutherford's experiment, a beam of alpha particles was scattered by atomic nuclei in a piece of gold foil. A lead shield protected researchers from the radiation.

Figure 21-1. The experiment started with a piece of radioactive material—matter that sends out energetic particles (see Chapter 26). For our purposes, you can think of radioactive materials as sources of tiny subatomic bullets. The particular material that Rutherford used produced bullets that scientists had named *alpha particles,* which are positively charged particles thousands of times heavier than electrons. By arranging the apparatus as shown, Rutherford produced a stream of these subatomic bullets moving toward the right in the figure. In front of this stream, he placed a thin foil of gold.

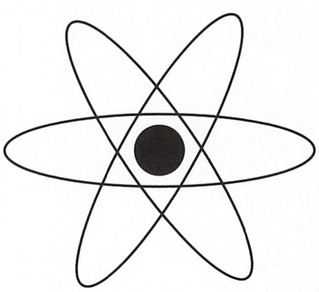
The experiment was designed to measure something about the way atoms are put together. At the time, people believed Thomson's idea that the small, negatively charged electrons were scattered around the entire atom, more or less like raisins in a bun. Rutherford was trying to shoot "bullets" into the "bun" to see what happened. He expected the results to confirm Thomson's model of the atom.

Instead, what the experiment revealed was little short of astonishing. Almost all the subatomic bullets either passed right through the gold foil unaffected or were scattered through very small angles. This result is easy to interpret: it means that most of the heavy alpha particles passed through spaces in between the gold atoms and that the alpha particles that hit the gold atoms were only moderately deflected by the relatively low-density material in them. However, about one alpha particle in a thousand was scattered through a large angle; some even bounced straight back. What could possibly cause the relatively heavy alpha particles to rebound in this way? Rutherford said that it was like firing a heavy artillery shell at a piece of tissue paper and having it bounce back and hit you.

After almost two years of puzzling over these extraordinary results, Rutherford concluded that a large part of each atom's mass is located in a very small, compact object at the center, which he called the **nucleus.** About 999 times out of 1000 the alpha particles either missed the atom completely or went through the low-density material in the outer regions of the atom. About 1 time out of 1000, however, the alpha particle hit the nucleus and bounced backward through a large angle.

You can think of the Rutherford experiment in this way. If the atom were a large ball of mist or vapor with a diameter greater than a skyscraper and the nucleus were a bowling ball at the center of that sphere of mist, then most bullets shot at the atom would go right through. Only those that hit the bowling ball would ricochet back through large angles. In this analogy, of course, the bowling ball plays the part of the nucleus, while the mist is the domain of the electrons.

As a result of Rutherford's work, a new picture of the atom emerged, one that is very familiar to us. Rutherford described a small, dense, positively charged nucleus sitting at the atom's center, with lightweight, negatively charged electrons circling it, like planets orbiting the Sun. (The nucleus must have a positive charge because the positively charged alpha particles bounced back from the nucleus rather than becoming attracted to it.) Indeed, Rutherford's discovery has become an icon of the modern age, adorning everyday objects from postage stamps to bathroom cleaners. Thus, by trying to confirm one model of the atom, Rutherford wound up proposing an entirely different model. This interplay of model and experiment has been typical of modern physics. In the first half of the twentieth century, many ideas about atomic structure were proposed, tested, and further refined. We examine many of these ideas in the next few chapters.



The Nuclear Regulatory Commission uses a highly stylized atomic model as its logo.

Later on, physicists discovered that the nucleus itself is made up primarily of two different kinds of particles (see Chapter 26). One of these carries a positive charge and is called a *proton*. The other, whose existence was not confirmed until 1932, carries no electric charge and is called a *neutron*.

For each positively charged proton in the nucleus of the atom, there is normally one negatively charged electron associated with the atom. The electric charges of the electrons and the protons are of equal magnitude and so cancel out; thus atoms are normally electrically neutral. In some cases, atoms either lose or gain electrons. In this case, they acquire an electric charge and are called *ions*.

Why the Rutherford Atom Couldn't Work

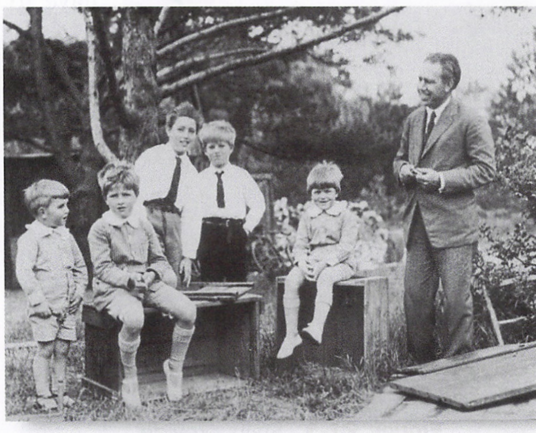
The picture of the atom that Rutherford developed is intellectually appealing, particularly because it recalls to us the familiar orbits of planets in our solar system. However, we have already learned enough about the behavior of nature to know that the atom just described could not possibly exist in nature. Why do we say this?

We learned in Chapter 3 that an object traveling in a circular orbit is constantly being accelerated—it is not in uniform motion because it is continually changing direction. Furthermore, we learned in Chapter 19 that any accelerated electric charge must give off electromagnetic radiation, as described by Maxwell's equations. Thus, if an atom really fit the Rutherford model, the electrons moving in their atomic orbits would constantly give off energy in the form of electromagnetic radiation. This energy, according to the first law of thermodynamics, would have to come from somewhere (remember conservation of energy!), so as the electrons gave up their energy to electromagnetic radiation, they would gradually spiral in toward the nucleus. Eventually, the electrons would fall into the nucleus and the atom would cease to exist in the form we know.

In fact, if you put in the numbers, the life expectancy of the Rutherford atom turns out to be less than a second. Given the fact that many atoms have survived billions of years, since almost the beginning of the universe, this calculation poses a serious problem for the simple orbital model of the atom.

WHEN MATTER MEETS LIGHT

Almost from the start, the Rutherford model of the atom encountered difficulties. Some of the problems involved its violations of fundamental physical laws as we have described, whereas others were more mundane—the Rutherford model simply did not explain all the behavior of atoms that scientists knew about. Rutherford and his contemporaries knew that the planetary model of the atom was a step forward from Dalton's idea of featureless spheres, but they also knew that, in the usual way of scientific progress, more work was needed to refine the model further. The first decades of the twentieth century were a period of tremendous ferment in physics as people scrambled to find a new way of describing the nature of atoms.



Niels Bohr (1885–1962) with his five sons, including Aage Bohr. Both won Nobel prizes in physics.

The Bohr Atom

In 1913, Niels Bohr (1885–1962), a young Danish physicist working in England, produced the first model of the atom that avoided the kinds of objections encountered by Rutherford's model. The **Bohr atom** does not match well with our intuition about the way things ought to be in the real world, but it was the precursor of the modern view of the atom's internal structure.

The young Bohr was deeply immersed in studying how atoms interact with light and other forms of electromagnetic radiation. He knew that some new ideas were circulating in theoretical physics at the time—ideas that things in the world of the atom were different from the way they were understood in the familiar Newtonian world. In particular, he knew that physicists were suggesting that in the atomic world, energy comes in discrete bundles, called *quanta* (see Chapter 22). Bohr wondered what the consequences would be if the angular momentum of electrons circling the nucleus also came in discrete units. In this case, electrons circling the nucleus, unlike planets circling the Sun, could not maintain their positions at just any distance from the center. Bohr found that if their angular momentum could only have certain discrete values, then there were only certain positions—he called them "allowed energy levels" or "allowed orbits"—located at specified distances from the center of the atom in which an electron could exist for long periods of time without giving off radiation. (As we shall see in Chapter 22, in the modern view of the atom, the analogy between electrons and planets is no longer accepted as completely accurate, although it provides a rough approximation of reality.)

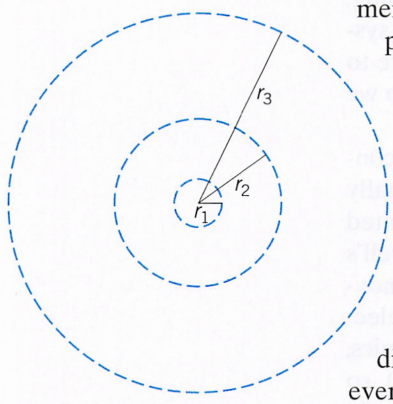


FIGURE 21-2. A schematic diagram of the Bohr atom showing the first three energy levels and respective distances (r_1, r_2, and r_3) from the nucleus.

Bohr's picture of the atom (Figure 21-2) embraces the idea that the electron can exist at a specific distance r_1 from the nucleus, at a distance r_2, or at a distance r_3, and so on, each distance corresponding to a different electron energy level. As long as the electron remains at one of those distances, its energy is fixed. In the Bohr atomic model, the electron cannot ever, at any time, exist at any place between these allowed energy states.

One way to think about the Bohr atom is to imagine what happens when you climb a flight of steps. You can stand on the first step or you can stand on the second step. It's very hard, however, to imagine what it would be like to stand somewhere between two steps. In just the same way, an electron can be in the first allowed energy level or in the second one, but it can't be in between these allowed energy levels. In terms of energy, both the steps and the electrons in an atom may be represented by a simple pictorial description (Figure 21-3). Each time you change steps in your home, your gravitational potential energy changes. Similarly, each time an electron changes levels, its energy changes.

An electron in an atom can be in any one of a number of allowed energy levels, each corresponding to a different distance from the nucleus. You would have to exert a force over a distance to move an electron from one allowed

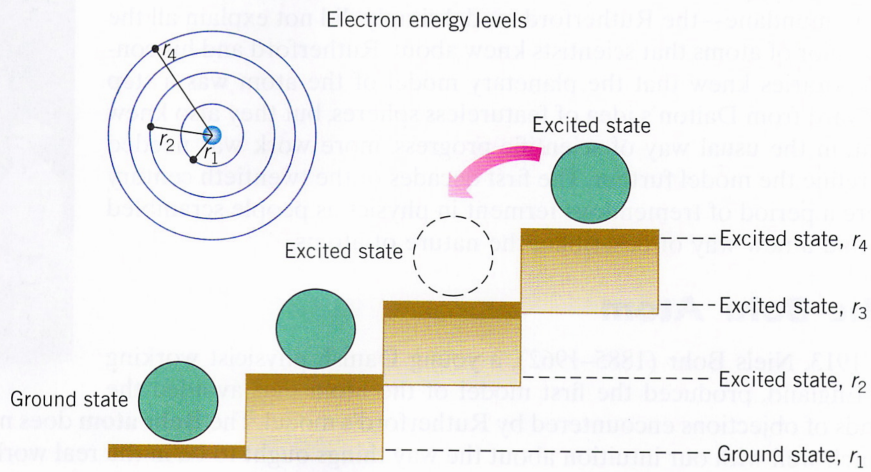


FIGURE 21-3. Stairs provide an analogy for the energy changes associated with electrons in the Bohr atom.

energy level to a higher level, just as your muscles have to exert a force to get you up a flight of stairs. Thus, the allowed energy levels of an atom occur as a series of steps, as shown in the figure. An electron in the lowest energy level is said to be in the *ground state,* while all energy levels above the ground state are called *excited states.*

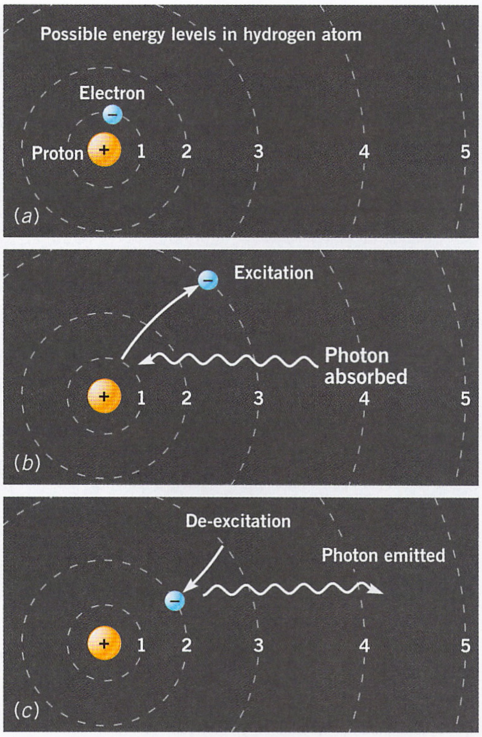


FIGURE 21-4. Electrons may jump between the energy levels shown in (*a*) and, in the process, (*b*) absorb or (*c*) emit energy in the form of a photon.

Photons: Particles of Light

One major feature of the Bohr atom is that an electron in a higher energy level can spontaneously move down into an available lower energy level. This process is analogous to that by which a ball at the top of a flight of stairs can bounce down the stairs under the influence of gravity.

Assume that an electron is in an excited state, as shown in Figure 21-4. The electron can move to a lower energy state, but if it does, something must happen to the extra energy. Energy can't just disappear. This realization was Bohr's great insight. The energy that's left over when the electrically charged electron moves from a higher state to a lower state is emitted by the atom in the form of a single packet of electromagnetic radiation—a particle-like bundle of light called a **photon.** Every time an electron jumps from a higher to a lower energy level, a photon moves away at the speed of light. The energy of the photon, which is proportional to its frequency, is equal to the difference between the electron's initial and final energy levels.

The concept of a photon raises a perplexing question: Is light—Maxwell's electromagnetic radiation—a wave or a particle? We explore this puzzle at some length in Chapter 22, once we have learned more about the behavior of atoms.

The interaction of atoms and electromagnetic radiation provides the most compelling evidence for the Bohr atom. If electrons are in excited states and if they make transitions to lower states, then photons are emitted. If we look at a group of atoms in which these transitions are occurring, we see light or other electromagnetic radiation. Thus, when you look at the flame of a fire or the fluorescent light in the ceiling, you are actually seeing photons that have been emitted by electrons jumping between allowed states in that material's atoms.

Not only does the Bohr model give us a picture of how matter emits radiation, it also provides an explanation for how matter absorbs radiation. For example, start with an electron in a low-energy state, perhaps its ground state. If a photon arrives that has just the right amount of energy to raise the electron to a higher allowed energy level (the next step up), the photon can be absorbed and the electron is pushed up to an excited state (see Figure 21-4*b*). The absorption of light is a mirror image of light emission.

This picture of the interaction of matter and radiation is exceedingly simple, but two key ideas are embedded in it. For one thing, when an electron moves from one allowed state to another, it cannot ever, at any time, be at any place in between. This rule is built into the definition of an allowed energy level. This means that the electron must somehow disappear from its original location and reappear in its final location without ever having to traverse any of the positions in between. This process, called a **quantum leap** or **quantum jump,** cannot be visualized, but it is something that seems to be fundamental in nature—an

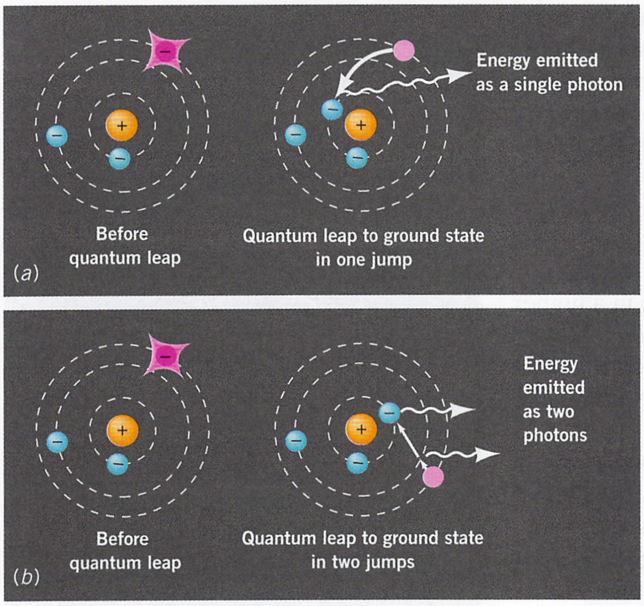
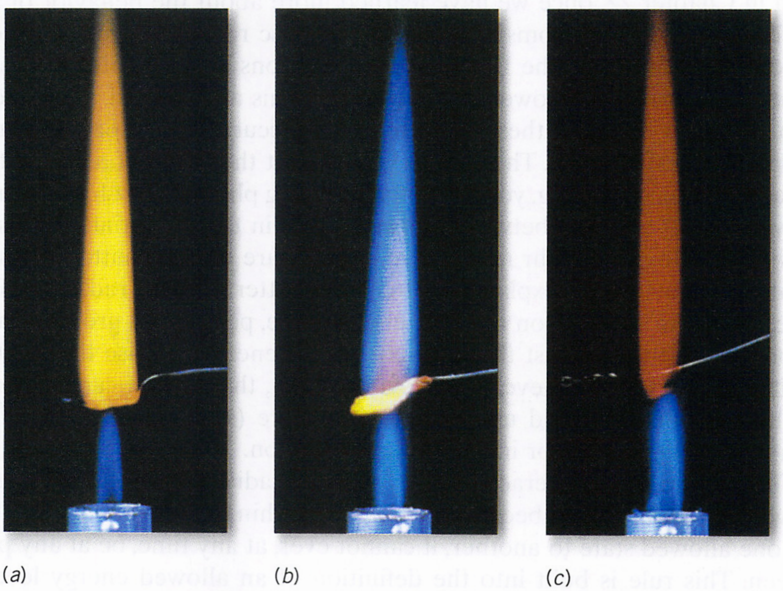


FIGURE 21-5. An electron can jump from a higher to a lower energy level in a single quantum leap (*a*) or by multiple quantum leaps (*b*).

example of the quantum weirdness of nature at the atomic scale that we discuss in Chapter 22.

The second key idea is that if an electron is in an excited state, it can, in principle, get back down to the ground state in one of several different ways. Look at Figure 21-5. An electron in the upper energy level can move to the ground state by making one large jump and emitting a single photon with a large amount of energy. Alternatively, it can move to the ground state by making two smaller jumps, as shown. Each of these smaller jumps emits a photon of somewhat less energy. The energies emitted in these smaller jumps are generally different from one another, but the sum of the two energies equals that of the single large jump. If we had a large collection of atoms of this kind, we would expect that some electrons would make the large leap while others would make the two smaller ones. Thus, when we look at a collection of these atoms, we would measure three different energies of photons.

This curious behavior of electron energy levels helps to explain the familiar phenomenon of fluorescence. Recall that the energy of electromagnetic radiation is related to its frequency. In fluorescence, the atom absorbs a higher-energy photon of ultraviolet radiation (which our eyes can't detect). The atom then emits two lower-energy photons, at least one of which is in the visible range. Consequently, shining ultraviolet black light on the fluorescent material, makes it glow with a bright, visible color.

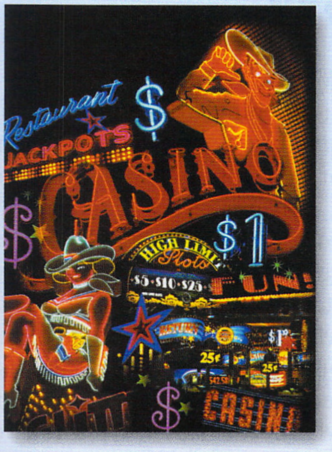


The elements (*a*) sodium, (*b*) potassium, and (*c*) lithium impart distinctive colors to a flame. The color from each element corresponds to the frequency of the photons emitted by that element as its electrons change energy level from higher energy to lower.

A key point about the Bohr model is that energy is required to lift an electron from the ground state to any excited state. This energy has to come from somewhere. We have already mentioned one possibility: that the atom absorbs a photon of just the right frequency to raise the electron to a higher energy level. There are other possibilities, however. For example, if the material is heated, atoms will move faster, gain kinetic energy, and undergo more energetic collisions. In these collisions, an atom can absorb energy and then use that energy to move electrons to a higher state. This explains why materials often glow when they are heated—the glow occurs when electrons drop back down after being raised to a higher energy level.

Develop Your Intuition: Bright Lights! Big City!

All big cities have large downtown areas of restaurants, stores, and movie theaters—all places with big electric signs in a dozen different colors. Most large electric light displays are gas discharge tubes—commonly called neon signs—in which gas at a low pressure glows when an electric current passes through it. How do these electric signs appear in so many different colors?



Bright neon lights decorate a city at night.

The glow of a gas discharge tube is caused by changes in the electron energy states in the gas atoms. Those atoms absorb energy, their electrons move to higher energy levels, and then the electrons fall back to lower energy levels and release the energy as light. The most common gas in these tubes is neon, which produces photons of orange-red light. If you want a different color, you can use a different gas with different energy levels. Chemically inert noble gases are particularly popular in gas discharge tubes: argon glows bluish-green, krypton glows bluish-white, and xenon glows blue. In fact, krypton emits such an intense light that it is commonly used for lighting airport runways. Other ways to produce bright colors with gas discharge tubes include mixing in other gases or coating the inside of the tube with fluorescent materials.

Gas discharge tubes are gradually being replaced with displays of light-emitting diodes (LEDs), which are solid crystals that also emit light when electrons fall back from excited energy levels. These remarkable materials are safer and they use less energy. However, the range of colors presently available is not nearly as wide as with gas discharge lights.

An Intuitive Leap

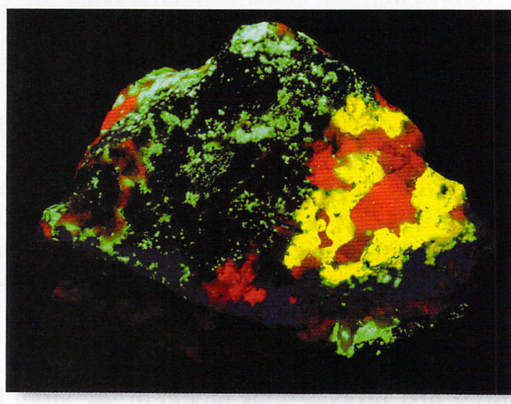
Bohr first proposed his model of the atom based on an intuition guided by experiments and ideas about the behavior of things in the subatomic world. In some ways, the Bohr model was completely unlike anything we experience in the macroscopic world; indeed, the model seemed to some a little bit crazy. It took two decades for scientists to develop a theory called *quantum mechanics* that showed why electrons can exist only within Bohr's allowed energy states and not in between them. We discuss this justification for the Bohr atom in Chapter 22, but remember that the justification occurred long after the initial hypothesis. Physicists accepted the Bohr model because it worked—it explained what they saw in nature and allowed them to make predictions about the behavior of real matter.

How could Bohr have come up with such a strange picture of the atom? He was, as we have said, guided by some of the early work that led to the theory of quantum mechanics (see Chapter 22). In the end, however, this explanation is unsatisfactory. Many people at the time studied interactions of atoms and light, but only Bohr was able to make the leap of intuition to his description of the atom. This insight, like Newton's realization that gravity might extend to the orbit of the Moon, remains one of the great intuitive achievements of the human mind.

SPECTROSCOPY

Whenever energy is added to a system with many atoms in it, electrons in some atoms jump to excited states. As time goes by, some of these electrons make quantum leaps down to the ground state, giving off photons as they do. If some of those photons are in the range of visible light, the source appears to glow.

You may not realize it, but you have looked at such collections of atoms all your life. Common mercury vapor streetlamps contain bulbs filled with mercury gas. When electric current passes through the gas, electrons move up to excited states. When they jump down, they emit photons that give the lamp a bluish-white color. Other types of streetlights, often used at freeway interchanges, use bulbs filled with sodium atoms. When sodium is excited, the most frequently emitted photons lie in the yellow range, so the lamps look yellow. Yet another place where you can see photons emitted directly by quantum leaps, as mentioned above, is in the vivid glowing colors of fluorescent objects, which are often used in black light displays at the theater.

Fluorescent minerals appear dull and ordinary under daylight, but glow with brilliant colors under ultraviolet light.

From these examples, we can draw two conclusions: (1) quantum leaps are very much in evidence in your everyday life, and (2) different atoms give off different characteristic photons. The second of these two facts is extremely important for physicists. If you think about the structure of an atom, the idea that different atoms emit and absorb different characteristic photons shouldn't be too surprising. Electron energy levels depend on the electrical attraction between the nucleus and the electrons, just as the orbits of the planets depend on the gravitational attraction between the planets and the Sun. Different nuclei have different numbers of protons, so electrons circling them are in different energy levels. In fact, the energies between allowed energy levels within atoms are different in each of the hundred or so different chemical elements. Because the energy and frequency of photons emitted by an atom depend on the differences in energy between these levels, each chemical element emits a distinct set of characteristic photons.

You can think of the collection of characteristic photons emitted by each chemical element as a kind of fingerprint—a set of wavelengths that is distinctive for that chemical element and none other. This feature opens up a very important application. The total collection of photons emitted by a given atom is called its **spectrum,** a characteristic pattern that can be used to identify chemical elements even when they are very difficult to identify by any other means. (Molecules also have individual spectra, but these are far more complicated than atomic spectra.)

In practice, the identification process works because light from the gaseous atoms is spread out after being passed through a prism (Figure 21-6). Each possible quantum jump corresponds to light at a specific wavelength, so each type

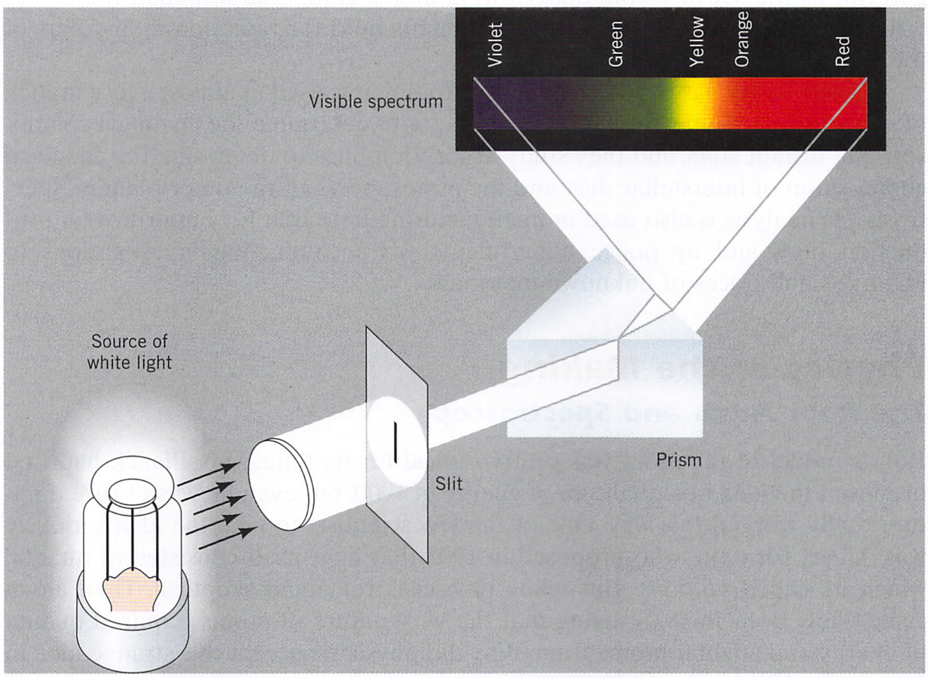


Figure 21-6. A glass prism spreads out the colors of the visible spectrum.

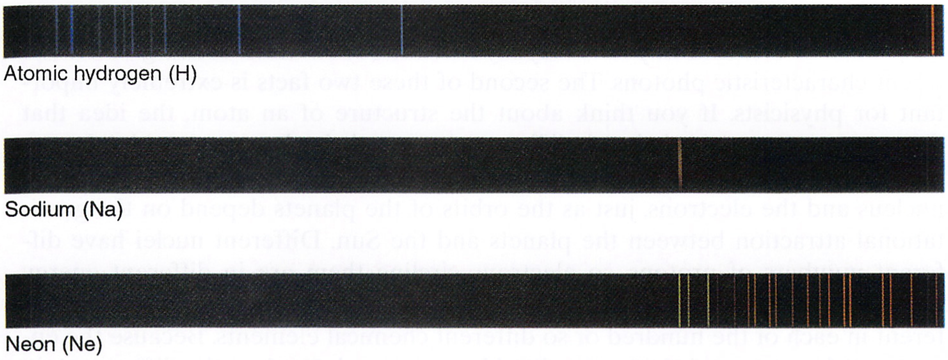


FIGURE 21-7. Line spectra, shown here for hydrogen, sodium, and neon, provide distinctive fingerprints for elements and compounds.

of atom produces a different set of lines, as shown in Figure 21-7. This spectrum is the atomic fingerprint.

The Bohr model suggests that if an atom gives off light of a specific wavelength and energy, then it also absorbs light at that wavelength. The emission and absorption processes, after all, involve quantum jumps between the same two energy levels, but in different directions. Thus, if white light shines through a material containing a particular kind of atom, then certain wavelengths of light are absorbed. When you observe that light on the other side of the material, then certain lines of color are missing. The dark areas corresponding to the absorbed wavelengths are called *absorption lines.* This set of lines is as much an atomic fingerprint as the set of colors that glowing atoms emit. And although the use of visible light is very common, these arguments hold for radiation in any part of the electromagnetic spectrum.

Spectroscopy has become a standard tool that is used in almost every branch of science. Astronomers use emission spectra to determine the chemical composition of distant stars, and they study absorption lines to determine the chemical composition of interstellar dust and the atmospheres of the outer planets. Spectroscopic analysis is also used in manufacturing to search for impurities on production lines and by police departments when conducting investigations to identify small traces of unknown materials.

Physics in the Making

The Bohr Atom and Spectroscopy

Bohr's model of the atom was pretty radical for its time. Max Planck had first proposed the idea of a quantum of energy in 1900, but even he wasn't sure if nature really worked this way. One of the few scientists to take the idea seriously was Albert Einstein, who proposed in 1905 that light itself consisted of quanta, which he called *photons.* But many physicists remained skeptical. Then along came Niels Bohr in 1913 saying that the very nature of atoms required quanta of energy and angular momentum. Why did physicists accept this strange idea in fewer than 10 years?

One of the great mysteries of the time, one of the mysteries alluded to at the beginning of this chapter, was the reason for spectroscopic patterns. Why should hydrogen, for example, emit light of just those wavelengths observed in its spectrum? Many scientists tried to find a formula for the pattern of wavelengths, without any luck. Finally, in 1885, a Swiss schoolteacher named Johann Balmer found a formula by trial and error that fit the wavelengths of the most prominent series of lines in hydrogen. Nobody knew why the formula worked, but it did.

Others got into the act. Hydrogen shows several series of lines in different parts of the electromagnetic spectrum, and each series is named after the person who discovered it: the Lyman series, the Paschen series, and so on. In 1890, the Swedish spectroscopist Johannes Rydberg found a formula that fit all these series. He had only to plug in simple integers such as 1, 2, and 3 to get Balmer's series, Lyman's series, or any other series. Rydberg's formula involved multiplying by a constant (now called "Rydberg's constant") with a value of 1.097×10^7 meter^{-1}. He didn't know why his constant had this value or why the formula worked, but it did.

When Bohr proposed his model of the atom, it explained how spectral lines are produced in the atom. He was able to derive a formula for the wavelengths emitted when electrons dropped from several excited levels in hydrogen to the ground state; the result matched Lyman's series. Bohr also derived a formula for the wavelengths emitted when electrons dropped from the excited levels in hydrogen to the first excited level; the results matched Balmer's series. All the series known for hydrogen turned out to fit particular transitions of electrons between allowed levels.

In all these formulas, Bohr had to multiply energies by a constant, but in his case it was a complicated factor involving the speed of light, the mass and charge of an electron, and other known constants. When Bohr plugged in these values and came up with a number for his constant, lo and behold, he got 1.097×10^7 meter^{-1}. Thus, Bohr showed how to calculate the Rydberg constant from more fundamental physical constants. This strong, direct confirmation of his theory went a long way toward winning acceptance of the Bohr model. (Bohr received the Nobel Prize for his work in 1922, one year after Einstein and four years after Planck.) ●

Connection

Spectra of Life's Chemical Reactions

In a classic set of experiments in the early 1940s, scientists used spectroscopy to work out in detail how large molecules called enzymes govern chemical reactions in living cells. In these experiments, a fluid containing the materials undergoing the chemical reactions was allowed to flow down a tube. As the fluid moved farther down the tube, the reaction progressed closer and closer to completion. By measuring spectra at different points along the tube, scientists were able to follow the changes in the behavior of electrons as the chemical reactions went along. In this way, part of the enormously complex problem of understanding the chemistry of life was unraveled.

More recently, scientists have begun to develop instruments that can use the principles of spectroscopy to identify pollutants emitted by automobile tailpipes as cars drive by. If they are successful, we will have a major new tool in our battle against air pollution and acid rain. ●

Physics in the Making

The Story of Helium

You have probably encountered helium, perhaps to inflate party balloons. Helium gas turns out to be a very interesting material, not only for its properties (it's less dense than air, so helium-filled balloons float up), but also because of the history of its discovery.

The word helium comes from "helios," the Greek word for Sun, because helium was first discovered by identifying a new set of spectral lines in light from the Sun, work done in 1868 by English scientist Joseph Norman Lockyer (1836–1920). Helium is very rare in Earth's atmosphere and before Lockyer's discovery scientists were not even aware of its existence. Following the discovery, there was a period of about 30 years when astronomers accepted the fact that the element helium existed in the Sun, but they were unable to find it on Earth.

The discovery of this hitherto unknown spectrum led to a very interesting problem. Could it be that there were chemical elements on the Sun that simply did not exist on our own planet? If so, it would call into question our ability to understand the rest of the universe, for the simple reason that if we don't know what an element is and can't isolate it in our laboratories, then we can never really be sure that we understand its properties. In fact, the existence of helium on Earth wasn't confirmed until 1895, when Sir William Ramsay identified its spectrum in a sample of radioactive material. ●

Connection

The Laser

2

The Bohr model provides an excellent way of understanding the workings of one of the most important devices in modern science and industry—the *laser.* The word "laser" is an acronym for *l*ight *a*mplification by *s*timulated *e*mission of *r*adiation. At the core of every laser is a collection of atoms—a crystal of ruby, perhaps, or a gas enclosed in a glass tube. The term "stimulated emission" refers to a process that goes on when light interacts with these atoms in a special way.

If an electron is in an excited state and one photon of just the right energy passes nearby, then the electron may be stimulated to make the jump to a lower energy state, thus releasing a second photon. By "just the right energy" for the first photon, we mean a photon whose energy corresponds to the energy gap between two electron energy levels in the atom.

What's so special about the photons emitted by the stimulated electron? Remember that light is a form of electromagnetic radiation that can be described as a wave. In a laser, the crests of all the emitted photon waves line up exactly with the crests of the first photon, and the signal is enhanced by constructive interference. In the language of physics, we say that the photons are "coherent." Thus, in stimulated emission you have one photon at the beginning of the process and two coherent photons at the end.

Now suppose that you have a collection of atoms where most of the electrons are in the excited state, as shown in Figure 21-8. If a single photon of the correct frequency enters this system from the left and moves to the right, it passes the first atom and stimulates the emission of a second photon. You then have two photons moving to the right. As these photons encounter other atoms, they, too, stimulate emission so that you have four photons. It's not hard to see that

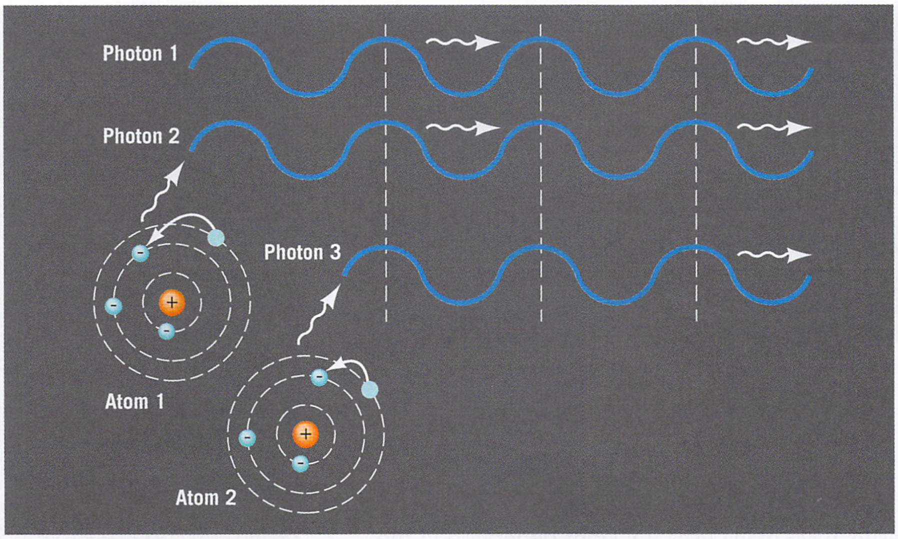


FIGURE 21-8. Lasers produce a beam of light when one photon stimulates the emission of other photons.

light amplification in a laser happens very quickly, cascading so that soon there is a flood of photons moving to the right through the collection of atoms. Energy added to the system from outside continuously returns atoms to their excited state—a process called "optical pumping"—so that more and more coherent photons can be produced.

In a laser (Figure 21-9), the collection of excited atoms is bounded on two sides by mirrors so that photons moving to the right hit the mirror, are reflected, and make another pass through the material, stimulating even more emissions of photons as they go. If a photon happens to be lined up exactly perpendicular to the mirrors at the end of the laser, it will continue bouncing back and forth. If its direction is off by even a small angle, however, it will eventually bounce out through the sides of the laser and be lost. Thus only those photons that are exactly aligned wind up bouncing back and forth between the mirrors, constantly amplifying the signal. Aligned photons traverse the laser millions of times, building up an enormous cascade of coherent photons in the system. Because only photons moving in exactly the right direction are amplified, the laser beam does not spread out very much, but stays tightly bunched. In the language of physics, we say that it is "collimated."

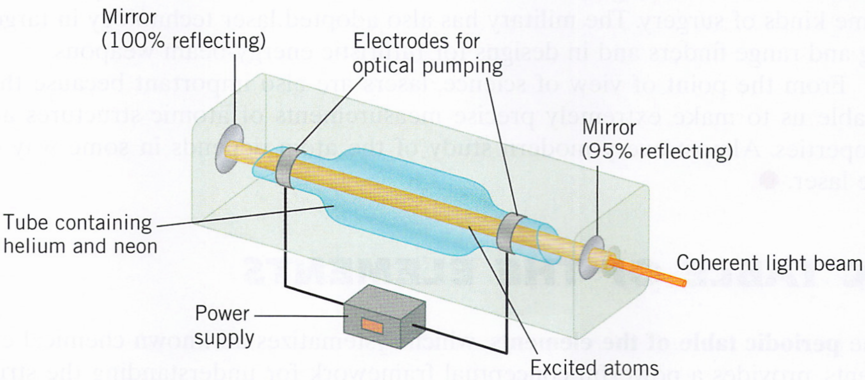
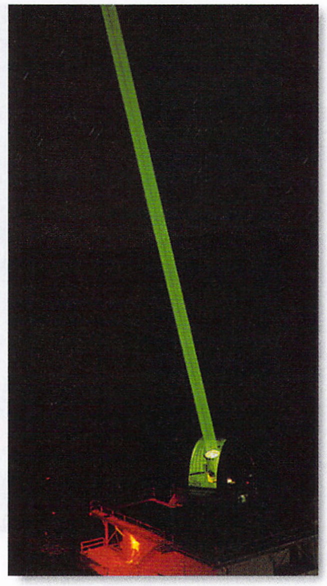


FIGURE 21-9. The action of a laser. Electrons in the laser's atoms are continuously "pumped" into an excited state by an outside energy source, and the beam of coherent photons is released when the electrons return to their ground state.

(*a*)



Lasers have many uses. (*a*) Light generated in the laser is carried by fiber-optic cable for use in surgery on the human ear. (*b*) Astronomers use lasers to align and focus telescopes.

(*b*)

The mirror from which the beam reflects is designed to be partially reflective—perhaps 95% of the photons that hit the mirror are reflected back into the laser. The remaining 5% of photons that leak out form the familiar laser beam, while the mirror at the other end reflects all the light that strikes it. Thus, the beam is made of intense, coherent light.

Laser beams have been applied in thousands of ways in science and industry since their development in the 1960s. Low-power lasers are ideal for optical scanners, such as the ones in supermarket checkout lines, and they make ideal light pointers for lectures and slide shows. The fact that the beam of light travels in a straight line makes the laser invaluable in surveying over long distances—for example, modern subway tunnels are routinely surveyed by using lasers to provide a straight line underground. Lasers are also used to detect the movement of seismic faults in order to predict earthquakes. In this case, a laser is directed across the fault, so that small motions of the ground are easily measured. Finely focused laser beams have revolutionized delicate procedures such as eye surgery. Much more powerful lasers that can transfer large amounts of energy are often used as cutting tools in factories, as well as implements for performing some kinds of surgery. The military has also adopted laser technology in targeting and range finders and in designs for futuristic energy-beam weapons.

From the point of view of science, lasers are also important because they enable us to make extremely precise measurements of atomic structures and properties. Almost every modern study of the atom depends in some way on the laser. ●

THE PERIODIC TABLE OF THE ELEMENTS

The **periodic table of the elements,** which systematizes all known chemical elements, provides a powerful conceptual framework for understanding the structure and interaction of atoms. Dmitri Mendeleev, the Russian scientist who developed the periodic table in the nineteenth century (see Chapter 1), assigned each element an integer called the *atomic number*. We now know that the atomic number corresponds to the number of protons in the atom, or, equivalently, if the atom is not charged, to the number of electrons surrounding the nucleus. If you arrange the elements as shown in Appendix C, with elements getting progressively heavier as you read from left to right and from top to bottom, as in a book, then elements in the same vertical column have very similar chemical properties.

Periodic Chemical Properties

The most striking characteristic of the periodic table is the similarity of the elements in any given column. For example, the far left-hand column of the table lists highly reactive elements called *alkali metals* (lithium, sodium, potassium, etc.). Each of these soft, silvery elements forms compounds (called salts) by combining in a one-to-one ratio with any of the elements in the seventh column (fluorine, chlorine, bromine, etc.). Water dissolves these compounds, which include sodium chloride, or table salt.

The elements in the second column (beryllium, magnesium, calcium, etc.) are metallic elements called the *alkaline earths,* and they too display similar chemical properties. For instance, these elements combine with oxygen in a one-to-one ratio to form colorless compounds with very high melting temperatures.

Elements in the far right-hand column (helium, neon, argon, etc.), by contrast, are all colorless, odorless gases that are almost impossible to coax into any kind of chemical reaction. These are called *noble gases* and they find applications when ordinary gases are too reactive. For example, helium is used to lift blimps because the only other lighter-than-air gas is the dangerous, explosive element hydrogen. Argon fills incandescent light bulbs because nitrogen or oxygen would react with the hot filament.

In the late nineteenth century, scientists knew that the periodic table worked—it organized the 63 elements known at that time and implied the existence of others—but they had no idea why it worked. Their faith in the periodic table was buttressed by the fact that, when Mendeleev first wrote it down, there were holes in the table—places where he predicted elements should go, but for which no element was known. The ensuing search for the missing kinds of atoms produced the elements we now call scandium (in 1876) and germanium (in 1886).

Why the Periodic Table Works: Electron Shells

With the advent of Bohr's atomic model and its modern descendants, we finally have some understanding of why the periodic table works. We now realize that the pattern of elements in the periodic table mirrors the spatial arrangement of electrons around the atom's nucleus—a concentric arrangement of electrons into *shells.*

The atom is largely empty space. When two atoms come near enough to one another to undergo a chemical reaction—such as a carbon atom and an oxygen atom in a burning piece of coal—electrons in the outermost shells meet one another first. These outermost electrons govern the chemical properties of materials, so we have to document the behavior of these electrons if we want to understand the periodic table.

In the process of studying atoms and their electrons, scientists discovered a curious fact. Electrons obey what is called the *Pauli exclusion principle,* which says that no two electrons in an atom can occupy the same state at the same time. (Note that the word "state" as used in this principle is not the same as an orbit or shell—there are normally many states in each shell.) One analogy is to compare electrons to cars in a parking lot. Each car takes up one space and once a space is filled, no other car can go there. Electrons behave just the same way. Once an electron fills a particular niche in the atom, no other electron can occupy the same niche. A parking lot can be full long before all the actual space in the lot is taken up with cars, because the driveways and spaces between cars must remain empty. So, too, a given electron shell can be filled with electrons long before all the available space is filled.

In fact, it turns out that there are only two spaces that an electron can fill in the innermost electron shell, which corresponds to the lowest Bohr energy level. One of these spaces corresponds to a situation in which the electron spins clockwise on its axis, the other to a situation in which it spins counterclockwise. When we start to catalog all possible chemical elements in the periodic table, we have element 1 (hydrogen) with a single electron in the innermost shell, and element 2 (helium) with two electrons in that same shell. After these two elements, if we want to add one more electron, it has to go into the second electron shell because the first electron shell is completely filled. (Note that this second shell is at a different energy than the first.) This situation explains why only hydrogen and helium appear in the first row in the periodic table.

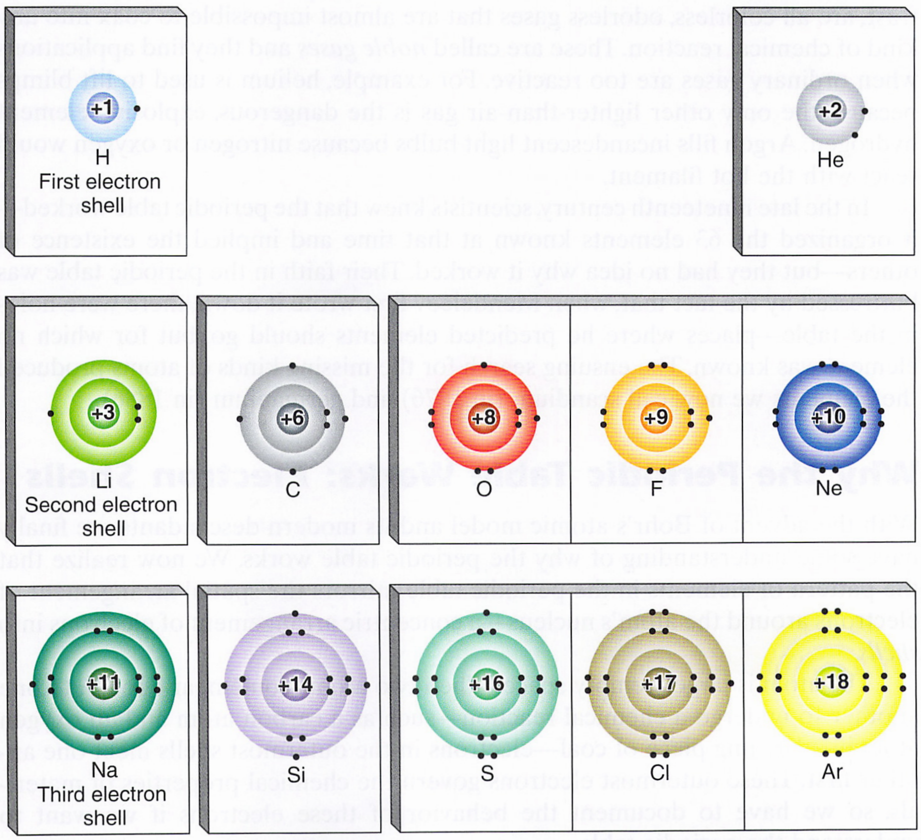


FIGURE 21-10. A representation of electrons in a number of common atoms.

Adding a third electron yields lithium, which is an atom with two electrons in the first shell and a single electron in the second electron shell. Lithium is the element just below hydrogen in the first column of the periodic table, because both hydrogen and lithium have a lone electron in their outermost shell.

The second electron shell has room for eight electrons, a fact reflected in the eight elements of the periodic table's second row, from lithium with three electrons to neon with ten. Neon appears directly under helium, and we expect these two elements to have similar chemical properties because both have a completely filled outer electron shell. In fact, both helium and neon are colorless, odorless, nonreactive gases.

Thus, a simple counting of the positions available to electrons in the first two electron shells explains why the first row in the periodic table has two elements in it and the second row has eight. By similar (but somewhat more complicated) arguments, you can show that the Pauli exclusion principle requires that the next row of the periodic table have 8 elements, the next 18, and so on (Figure 21-10). Thus, with an understanding of the shell-like structure of the atom's electrons, the mysterious regularity that Mendeleev found among the chemical elements becomes an example of nature's laws at work.

Develop Your Intuition: Predicting Chemical Formulas

The periodic table tells you how many electrons are in the outer shell of each element. We know that atoms are usually most stable (less likely to react chemically with other atoms) with a completely filled outer shell of electrons, which means eight electrons in most cases. Can you use that information to predict the formulas of simple compounds of the elements?

Elements at the left in the periodic table generally donate electrons to form a stable outer shell. Elements at the right in the periodic table generally accept electrons to complete a stable outer shell. Thus, an atom with one electron in its outer shell can combine with an atom with seven electrons in its outer shell, forming a compound with elements in a one-to-one ratio. Table salt, sodium chloride, has one atom of sodium for every atom of chlorine. Other alkali metals form similar compounds with halogens, such as potassium iodide or lithium fluoride. Atoms with two electrons in the outer shell can combine in one-to-one ratio with atoms containing six electrons in the outer shell. Examples are calcium oxide and magnesium sulfide.

Atoms can also form molecules containing three atoms, where two atoms containing one electron in the outer shell combine with an atom containing six electrons in the outer shell. This way you get potassium oxide, with two atoms of potassium for every atom of oxygen, and sodium sulfide, with two atoms of sodium for one of sulfur. You can also have molecules formed from one atom with two electrons in its outer shell and two atoms with seven electrons in the outer shell. Examples are calcium chloride, with two atoms of chlorine for every one of calcium, and magnesium fluoride, with two atoms of fluorine for every atom of magnesium.

See if you can figure out other combinations of atoms that might form molecules. How many atoms of sodium combine with one atom of nitrogen? How many atoms of aluminum combine with how many atoms of oxygen?

THINKING MORE ABOUT

Atoms: What Do Atoms Look Like?

Throughout this book you will find drawings of atoms. In this chapter we draw atoms as electrons in circular shells around a central nucleus. In Dalton's original work, atoms appear as little spheres in pictures of molecules. In other chapters in this book, atoms are portrayed as fuzzy clouds, waves, or even collections of dozens of smaller spherelike particles. So, what do atoms really look like?

Strictly speaking, we only see something when electromagnetic waves from the visible part of the spectrum enter our eyes. We are accustomed, however, to talking about other ways of "seeing." You cannot see X rays being absorbed by your teeth unless some intermediary system—a film or an electronic video monitor, for example—converts the X rays into a pattern that can be detected in the visible region of the electromagnetic spectrum. Similarly, astronomers often convert data from radio waves, infrared radiation, and other radiation into false-color images of distant objects. Scanning tunneling microscope pictures of atoms come from another such transformation—the amount of electric charge at a particular point on a material's surface is converted by electronics into peaks and valleys on a digital image.

A scanning tunneling microscope image of individual iron atoms on copper forms the Japanese character for "atoms."

So what does an atom look like? Is an X-ray picture of your teeth more real than the scanning tunneling microscope picture of the atom? Why or why not? Does the fact that we can't see atoms with our eyes mean that they don't exist?

Summary

All the solids, liquids, and gases around us are composed of about 100 different elements. Atoms, the building blocks of our chemical world, combine into groups of two or more; these groups are called molecules. Each atom contains a massive central **nucleus** made from positively charged protons and electrically neutral neutrons. Surrounding the nucleus are **electrons,** which are negatively charged particles that have only a small fraction of the mass of protons and neutrons.

Early models of the atom treated electrons like the planets orbiting around the Sun. Those models were flawed, however, because each electron, constantly accelerating, would have to emit electromagnetic radiation continuously. Niels Bohr proposed an alternative model in which electrons exist in various energy levels, much as you can stand on different levels of a flight of stairs.

Electrons in the **Bohr atom** can shift to a higher energy level by absorbing the energy of heat or light. Electrons can also drop into a lower energy level and in the process release heat or a **photon,** an individual electromagnetic wave. These changes in electron energy level are called **quantum leaps** or **quantum jumps. Spectroscopic** studies of the light emitted or absorbed by atoms—the atom's **spectrum**—reveal the nature of each atom's electron energy levels.

Each atom's electrons are arranged in concentric shells. When two atoms interact, electrons in the outermost shell come into contact. This shell-like electronic structure is reflected in the organization of the **periodic table of the elements,** which lists all the elements in rows corresponding to increasing numbers of electrons in each shell and in columns corresponding to elements with similar numbers of outer-shell electrons and thus similar chemical behavior.

Key Terms

Bohr atom A model of the atom proposed in 1913 by Danish physicist Niels Bohr; the Bohr atom revolutionized our understanding of physics. (p. 451)

electron A negatively charged fundamental particle; it is one of the primary building blocks of the atom. (p. 449)

nucleus A very small, dense, positively charged object at the center of every atom; nuclei are made up of protons and neutrons. (p. 450)

periodic table of the elements The systematic organization of the known chemical elements in terms of their electron configurations. (p. 462)

photon The quantum of electromagnetic radiation; a particlelike bundle of light. (p. 453)

quantum leap or **quantum jump** The process of changing location without having to traverse any of the positions in between; usually in reference to electrons changing energy levels in an atom. (p. 453)

spectroscopy The analysis of emission and absorption spectra of materials to determine their chemical properties. (p. 458)

spectrum The total collection of photons emitted by an atom; the distribution of the frequencies of photons emitted by a radiating system. (p. 457)

Review

1. What is an element? How many elements exist in nature?
2. What are most stars made of?
3. What three particles make up almost every atom? What are the major differences among these particles?
4. Review the basic components of Rutherford's experiment. What was it about the results of the experiment that led Rutherford to the conclusion that the atom has a nucleus?
5. In what ways is the structure of an atom like the solar system of planets orbiting around the Sun? In what ways is it different?
6. How does the nucleus differ from electrons?
7. What makes atoms neutral with respect to electric charge?
8. What is an ion?
9. Think of an analogy for the Rutherford experiment other than the cloud-of-mist-plus-bowling-ball we describe in the chapter. How does your analogy differ from that one? In what ways is it better or worse?
10. Why would an atom like Rutherford's model constantly give off energy, and how would this affect the electron orbit?
11. How does Bohr's model of the atom differ from that of Rutherford's? What is an allowed energy level?
12. What are quanta? How did this concept influence Bohr's thinking on atomic structure?
13. How is the Bohr atom similar to a set of steps?
14. What is the ground state of an electron? An excited state?
15. What is a photon?
16. What are you actually seeing when you look at a fire or the red-hot coil of an electric stove?
17. Describe the relationship between a photon and a quantum leap. What makes such a leap or jump occur?
18. Describe the alternative ways an electron in an excited state can return to its ground state. How much energy is emitted in each case as the electron returns to this ground state, and in what form is it emitted?
19. What is fluorescence?
20. Cite three examples of everyday objects with bright emission spectra.
21. Do different atoms give off different characteristic photons when excited? What is it about the structures of different elements that would lead to an element emitting photons of unique energies?
22. What is the spectrum of an element?
23. How is the emission spectrum of an element detected? What does each individual line stand for?
24. What is an absorption spectrum, and what is the specific technique used to capture it? What do the lines of this spectrum stand for?
25. Compare and contrast an emission spectrum with an absorption spectrum.
26. How might astronomers on Earth use spectroscopy to determine chemical elements that occur in stars?
27. How was helium discovered?
28. Describe the basic components of a laser. How does a laser work?
29. What is meant by optical pumping in a laser, and what specific role do mirrors play in helping to generate laser light? What is the 5% of reflected light that leaks out?
30. What is the atomic number of an element?
31. How was the periodic table of elements developed? How did the holes in the original table eventually help confirm its validity?

32. What does it mean to say the periodic table was useful because it worked? How does this relate to the scientific method? Describe an imaginary discovery that might have invalidated the periodic table.

33. In what ways are all the elements in a given column of the periodic table similar? In what ways are they different?

34. What is an electron shell? How many electrons are in the first shell of an atom? The second?

35. What is the Pauli exclusion principle?

36. How do electron shells help explain how the periodic table works?

37. What do atoms really look like?

Questions

1. The leaves of a tree are bright green. What do you think a leaf's absorption spectrum might look like?

2. Draw three different emission spectra that would appear red. (*Hint:* Think about different kinds of red objects, including a red laser, a red-hot coal, and a red sweater.)

3. Rutherford's experiment involved firing nucleus-sized bullets at atoms of gold. He found that one atom in 1000 bounced backward. Using the kind of simple pictures of the atom introduced in the chapter, speculate about how the experiment might have turned out if atoms were completely uniform in mass. What if electrons were more massive than the nucleus? (*Hint:* What happens when a bowling ball collides with a Ping-Pong ball?)

4. Advertisers often describe improvements in their products as a "quantum leap." Is this an appropriate use of the term? How big is a quantum leap?

5. Based on your knowledge of Newton's laws of motion, the laws of thermodynamics, and the nature of electromagnetic radiation, explain why the Rutherford model of the atom couldn't work.

6. When you shine invisible ultraviolet light (black light) on certain objects, they glow with brilliant colors. How can this behavior be explained in terms of the Bohr atom?

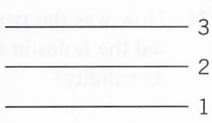
7. Why do different lasers have beams with different colors?

8. Space probes often carry compact spectrometers among their scientific hardware. What kind of spectroscopy might scientists use to determine the surface composition of the cold, outer planets that orbit the Sun? How might they use spectroscopy to determine the atmospheric composition of these planets?

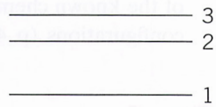
9. Suppose a particular atom has only two allowable electron orbits. How many different wavelength photons (spectral lines) would result from all electron transitions in this atom?

In the following three questions the energy levels of the atom are represented by horizontal lines, where the vertical spacings between the lines are proportional to the energy differences between the levels.

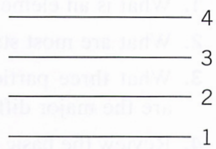
10. Suppose an atom has three equally spaced energy levels, as shown. How many different photons would result from all possible electron jumps between levels? Which jump corresponds to the highest frequency photon?



11. Suppose an atom has three energy levels, as shown. How many different photons would result from all possible electron jumps between levels? Which jump corresponds to the longest wavelength photon?

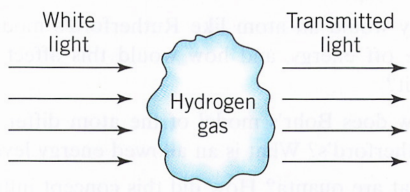


12. An atom has four equally spaced energy levels, as shown. How many different photons would result from all possible electron jumps between levels? Which jump(s) corresponds to the longest wavelength photon? Which jump corresponds to the photon with the highest frequency?



13. Figure 21-4 shows five of the energy levels in the hydrogen atom. Consider three electron jumps: from level 5 to level 2, from level 4 to level 2, and from level 3 to level 2. These jumps give off red, blue-green, and violet photons. Which jumps correspond to which colors? Explain.

14. A beam of white light shines through a sample of cool hydrogen gas, as shown. Describe the light that comes out the other side. The hydrogen is then heated to a very high temperature; describe the light that comes out the other side.



15. In his famous experiment, Rutherford fired alpha particles at a thin gold film. Most of the alpha particles went through the film and a very few bounced back. Suppose instead that about one-half the alpha particles bounced back and one-half went through. How would this have changed his conclusion about the structure of the atom?

16. In the process of fluorescence, an atom absorbs a photon of ultraviolet light and emits two or more photons of visible light. Is the reverse process possible? That is, it is possible for an atom to absorb a photon of visible light and emit photons of ultraviolet light?

17. A 100-watt lightbulb becomes warm and glows brightly enough to light a small room. On the other hand, a 100-watt laser can cut holes in steel and would not be effective at lighting a small room. What is it about the light coming from these two sources that accounts for these differences?

18. Silicon (Si) and nitrogen (N) are adjacent to carbon (C) on the periodic table. Si and C have many similar chemical properties but C and N do not. What accounts for this difference?

19. Explain why sodium chloride (NaCl) is such a stable compound.

20. Explain why two hydrogen (H) atoms combine with one oxygen (O) atom to form water (H_2O).

Problems

1. How many protons do the following elements have?

a. Hydrogen (H)
b. Carbon (C)
c. Sulfur (S)
d. Calcium (Ca)
e. Iodine (I)

2. How many electrons does each of the elements in Problem 1 have when they are electrically neutral?

3. How many electrons do the following ions have?

a. +1 of sodium (Na)
b. −2 of oxygen (O)
c. +3 of iron (Fe)
d. −1 of bromine (Br)
e. +2 of calcium (Ca)
f. +4 of lead (Pb)

4. How many protons do the ions in Problem 3 have? Does this differ from the number these elements would have if they were electrically neutral?

5. If a one-electron atom can occupy any of four different energy levels, how many lines might appear in that atom's spectrum?

6. If you were told that fluorine is an extremely reactive element (that is, it combines readily with other elements), what other elements could you guess were also extremely reactive? Why?

7. If you were told that argon (Ar) is an exceptionally unreactive element, what other elements could you guess were also extremely unreactive? Why?

Investigations

1. Investigate the history of the discovery of the chemical elements. What technological innovations led to the discovery of several new elements? Which was the most recent element to be discovered and how was it found?

2. Simple handheld spectrometers are available in many science labs. Use one to look at the spectra of different kinds of lightbulbs: an incandescent bulb, a fluorescent bulb, a halogen bulb, and any other kinds available to you. What differences do you observe in their spectra? Why?

3. Place pieces of transparent materials between a strong light source and the spectrometer mentioned in Investigation 2. Does the spectrum change? Why?

4. Why do colors look different when viewed indoors under fluorescent light than outdoors in sunlight? How might you devise an experiment to quantify these differences?

5. Investigate the variety of lasers that are currently available. What is the range of wavelengths available? How are different lasers used in medicine? In industry? In science?

WWW Resources

See the *Physics Matters* home page at **www.wiley.com/college/trefil** for valuable web links.

1. http://www.chemistry.org/portal/a/c/s/1/acsdisplay.html?DOC=sitetools\periodic_table.html A sophisticated periodic table of the elements site by the American Chemical Society. See especially the filling of electron shells.

2. http://www.achilles.net/~jtalbot/history/ A large site devoted to the history of the laser, including rich links to laser physics applications, design, related phenomena, and newsworthy lasers.

3. http://www.achilles.net/~jtalbot/data/elements/index.html A site devoted to gas discharge spectra of the light elements.

4. **http://www.colorado.edu/physics/PhysicsInitiative/Physics2000/elements_as_atoms/index.html** An excellent site devoted to the quantum atom. This location within that site runs an animated tutorial with simulations that describes the physical properties of electrons within elements, followed by a tutorial on the periodic table.
5. **http://www.colorado.edu/physics/PhysicsInitiative/Physics2000/quantumzone/index.html** An excellent site devoted to the quantum atom. This location within that site runs an animated tutorial with simulations that starts with gas discharge spectra, leads through spectroscopy, the Bohr atom and electron energy levels.
6. **http://www.almaden.ibm.com/vis/stm/gallery.html** A scanning tunneling microscope image gallery from IBM. The art and beauty of atomic images.
7. **http://micro.magnet.fsu.edu/electromag/java/rutherford/index.html** A Java applet simulation of the Rutherford Experiment from Florida State University.
8. **http://www.aip.org/history/electron/** An online exhibit describing the discovery of the electron by J.J. Thompson from the American Institute of Physics.