CHAPTER 4
Arrangement of Electrons in Atoms

BIG IDEA
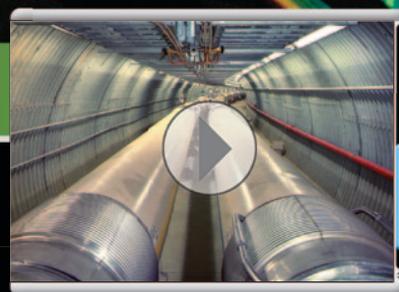An atom's nucleus is surrounded by electrons that have both a wave and a particle-like behavior.

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Atoms

The Development of a New Atomic Model

Key Terms

electromagnetic radiation
electromagnetic spectrum
wavelength
frequency

photoelectric effect
quantum
photon
ground state

excited state
line-emission spectrum
continuous spectrum

The Rutherford model of the atom was an improvement over previous models, but it was incomplete. It did not explain how the atom's negatively charged electrons are distributed in the space surrounding its positively charged nucleus. After all, it was well known that oppositely charged particles attract each other. So what prevented the negative electrons from being drawn into the positive nucleus?

In the early twentieth century, a new atomic model evolved as a result of investigations into the absorption and emission of light by matter. The studies revealed a relationship between light and an atom's electrons. This new understanding led directly to a revolutionary view of the nature of energy, matter, and atomic structure.

▶ **MAIN IDEA**

Light has characteristics of both particles and waves.

Before 1900, scientists thought light behaved solely as a wave. This belief changed when it was later discovered that light also has particle-like characteristics. Still, many of light's properties can be described in terms of waves. A quick review of these wavelike properties will help you understand the basic theory of light as it existed at the beginning of the twentieth century.

The Wave Description of Light

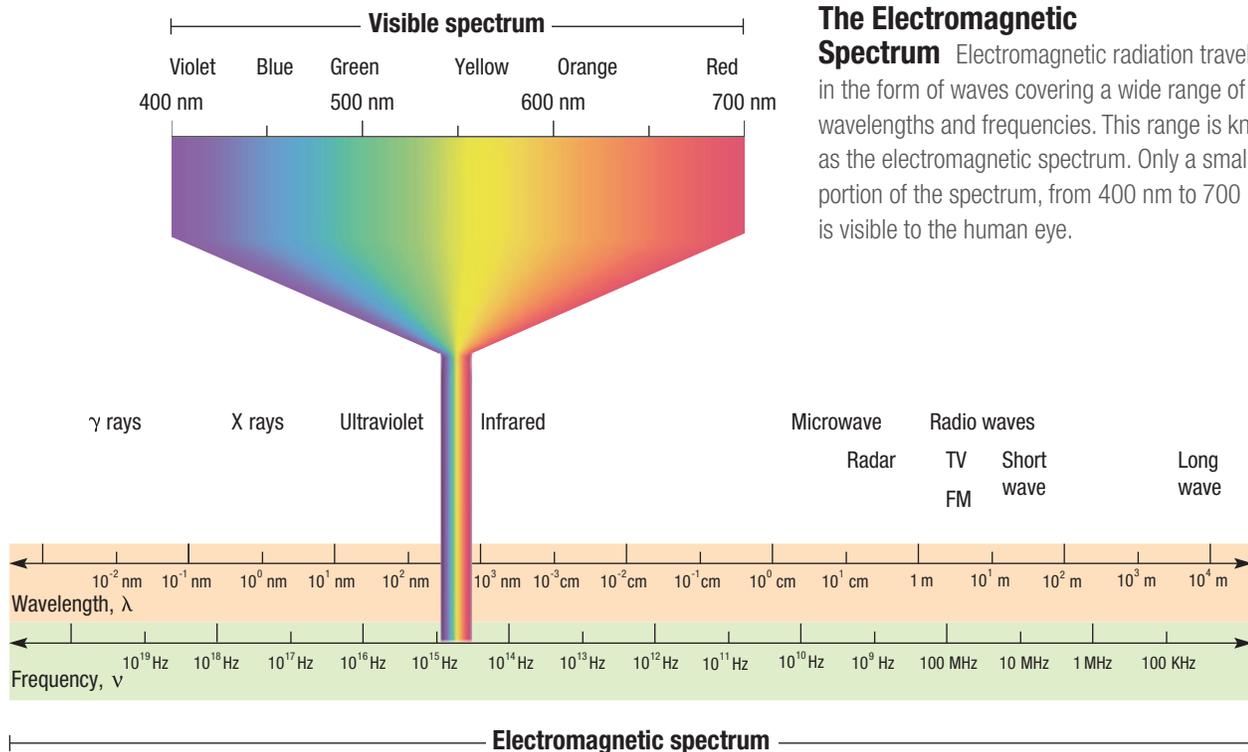**Visible light is a kind of electromagnetic radiation, which is a form of energy that exhibits wavelike behavior as it travels through space.** Other kinds of electromagnetic radiation include x-rays, ultraviolet and infrared light, microwaves, and radio waves. **Together, all the forms of electromagnetic radiation form the electromagnetic spectrum.** The electromagnetic spectrum is represented in **Figure 1.1** on the next page. All forms of electromagnetic radiation move at a constant speed of 3.00×10^8 meters per second (m/s) through a vacuum and at slightly slower speeds through matter. Because air is mostly empty space, the value of 3.00×10^8 m/s is also light's approximate speed through air.

The significant feature of wave motion is its repetitive nature, which can be characterized by the measurable properties of wavelength and frequency. **Wavelength (λ) is the distance between corresponding points on adjacent waves.** The unit for wavelength is a distance unit.

FIGURE 1.1

Visible spectrum

| Violet | Blue | Green | Yellow | Orange | Red |

400 nm   500 nm   600 nm   700 nm

The Electromagnetic Spectrum Electromagnetic radiation travels in the form of waves covering a wide range of wavelengths and frequencies. This range is known as the electromagnetic spectrum. Only a small portion of the spectrum, from 400 nm to 700 nm, is visible to the human eye.

γ rays   X rays   Ultraviolet   Infrared   Microwave   Radio waves

Radar   TV   Short wave   Long wave

FM

Wavelength, λ: 10^{-2} nm 10^{-1} nm 10^{0} nm 10^{1} nm 10^{2} nm 10^{3} nm 10^{-3} cm 10^{-2} cm 10^{-1} cm 10^{0} cm 10^{1} cm   1 m 10^{1} m 10^{2} m 10^{3} m 10^{4} m

Frequency, ν: 10^{19} Hz 10^{18} Hz 10^{17} Hz 10^{16} Hz 10^{15} Hz 10^{14} Hz 10^{13} Hz 10^{12} Hz 10^{11} Hz 10^{10} Hz 10^{9} Hz   100 MHz   10 MHz   1 MHz   100 KHz

Electromagnetic spectrum

Depending on the type of electromagnetic radiation, it may be expressed in meters, centimeters, or nanometers, as shown in **Figure 1.1.**

Frequency (ν) is defined as the number of waves that pass a given point in a specific time, usually one second. Frequency is expressed in waves/second. One wave/second is called a hertz (Hz), named for Heinrich Hertz, who was a pioneer in the study of electromagnetic radiation. **Figure 1.2** illustrates the properties of wavelength and frequency for a familiar kind of wave, a wave on the surface of water. Frequency and wavelength are mathematically related to each other.

FIGURE 1.2

Wavelength and Frequency

The distance between two corresponding points on one of these water waves, such as from crest to crest, is the wave's wavelength, λ. We can measure the wave's frequency, ν, by observing how often the water level rises and falls at a given point, such as at the post.

✔ **CRITICAL THINKING**

Analyze Which waves have a greater wavelength, those in (a) or those in (b)? Which have the greater frequency? Explain each answer.

To the beach ⟶

λ

(a)

λ

(b)
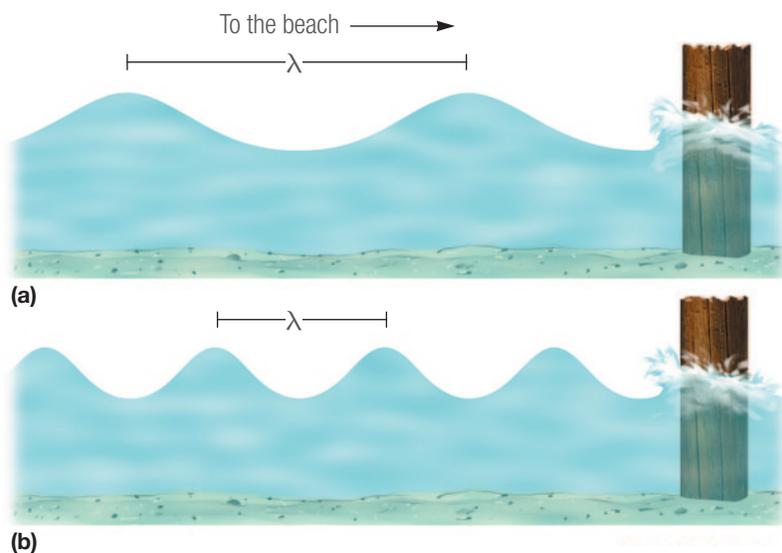
For electromagnetic radiation, the mathematical relationship between frequency and wavelength is written as follows:

$$c = \lambda\nu$$

In the equation, c is the speed of light (in m/s), λ is the wavelength of the electromagnetic wave (in m), and ν is the frequency of the electromagnetic wave (in s^{-1}). Because c is the same for all electromagnetic radiation, the product $\lambda\nu$ is a constant.

▶ MAIN IDEA

When certain frequencies of light strike a metal, electrons are emitted.

In the early 1900s, scientists conducted two experiments involving inter-actions of light and matter that could not be explained by the wave theory of light. One experiment involved a phenomenon known as the photo-electric effect. **The photoelectric effect refers to the emission of electrons from a metal when light shines on the metal,** as illustrated in **Figure 1.3**.

The mystery of the photoelectric effect involved the frequency of the light striking the metal. For a given metal, no electrons were emitted if the light's frequency was below a certain minimum—regardless of the light's intensity. Light was known to be a form of energy, capable of knocking loose an electron from a metal. But the wave theory of light predicted that light of any frequency could supply enough energy to eject an electron. Scientists couldn't explain why the light had to be of a minimum frequency in order for the photoelectric effect to occur.

The Particle Description of Light

The explanation of the photoelectric effect dates back to 1900, when German physicist Max Planck was studying the emission of light by hot objects. He proposed that a hot object does not emit electromagnetic energy continuously, as would be expected if the energy emitted were in the form of waves. Instead, Planck suggested that the object emits energy in small, specific packets called quanta. **A quantum of energy is the minimum quantity of energy that can be lost or gained by an atom.** Planck proposed the following relationship.

> **Quantum of Energy** $E = h\nu$

In the equation, E is the energy, in joules, of a quantum of radiation, ν is the frequency, in s^{-1}, of the radiation emitted, and h is a fundamental physical constant now known as Planck's constant; $h = 6.626 \times 10^{-34}$ J • s.

We can relate a quantum of energy to the speed of light as follows. Because $c = \lambda\nu$, we know that $\nu = c/\lambda$. Substituting this value for ν into the equation for a quantum of energy gives the following: $E = h\nu$, and $\nu = c/\lambda$, so $E = hc/\lambda$.
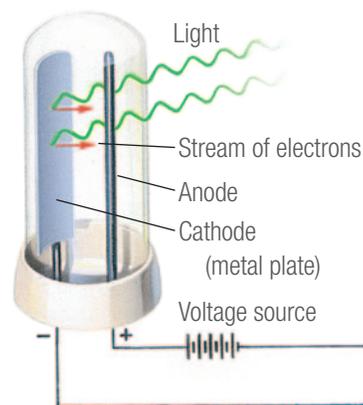
For example, for light of wavelength $\lambda = 500.0$ nm,

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\,\text{J} \bullet \text{s})(3.00 \times 10^{8}\,\text{m/s})}{5.000 \times 10^{-7}\,\text{m}} = 3.98 \times 10^{-19}\,\text{J}.$$

FIGURE 1.3

The Photoelectric Effect Electromagnetic radiation strikes the surface of the metal, ejecting electrons from the metal and causing an electric current.

Light

Stream of electrons

Anode

Cathode (metal plate)

Voltage source

In 1905, Albert Einstein expanded on Planck's theory by introducing the radical idea that electromagnetic radiation has a dual wave-particle nature. Light exhibits many wavelike properties, but it can also be thought of as a stream of particles. Each particle carries a quantum of energy. Einstein called these particles photons. **A photon is a particle of electromagnetic radiation having zero mass and carrying a quantum of energy.** The energy of a photon depends on the frequency of the radiation, or $E_{photon} = h\nu$.

Einstein explained the photoelectric effect by proposing that electromagnetic radiation is absorbed by matter only in whole numbers of photons. For an electron to be ejected from a metal surface, the electron must be struck by a single photon possessing at least the minimum energy required, which corresponds to a minimum frequency. Electrons in different metals are bound more or less tightly, so different metals require different minimum frequencies to exhibit the photoelectric effect.

▶ MAIN IDEA

Electrons exist only in very specific energy states for atoms of each element.

When current is passed through a gas at low pressure, the potential energy of the gas atoms increases. **The lowest energy state of an atom is its ground state. A state in which an atom has a higher potential energy than it has in its ground state is an excited state.** There are many possible excited states, each with a unique energy, but only one ground-state energy for atoms of a given element. When an excited atom returns to its ground state or a lower-energy excited state, it gives off the energy it gained in the form of electromagnetic radiation. The production of colored light in neon signs, as shown in **Figure 1.4**, is a familiar example of this process.

FIGURE 1.4

Light Emission Excited neon atoms emit light when electrons in higher energy levels fall back to the ground state or to a lower-energy excited state.

FIGURE 1.5

Emission-Line Spectra Excited hydrogen atoms emit a pinkish glow, as is shown in this diagram. When the visible portion of the emitted light is passed through a prism, it is separated into specific wavelengths that are part of hydrogen's emission-line spectrum. The line at 397 nm is in the ultraviolet and is not visible to the human eye.
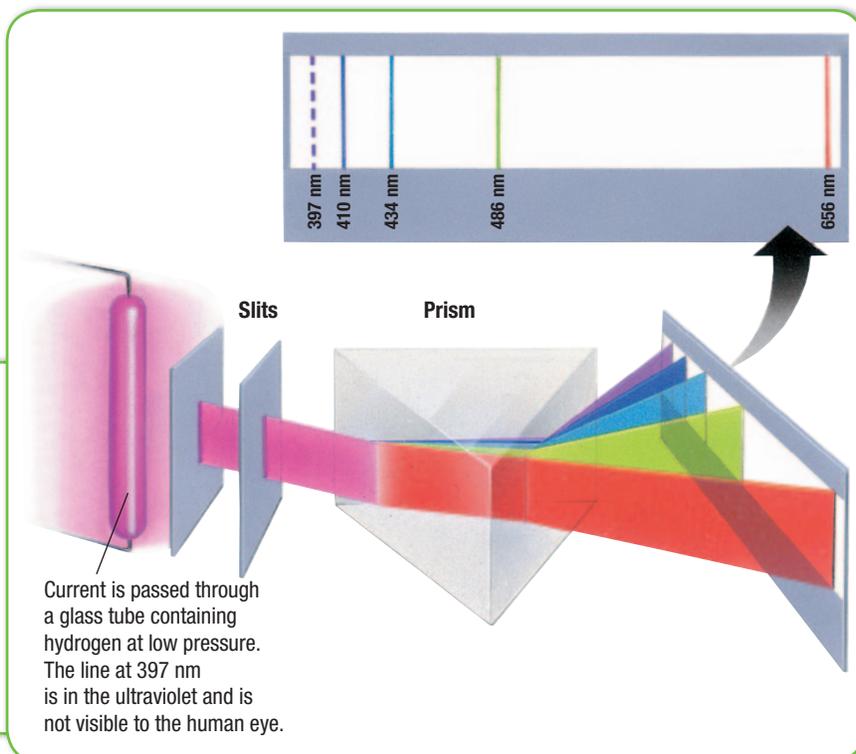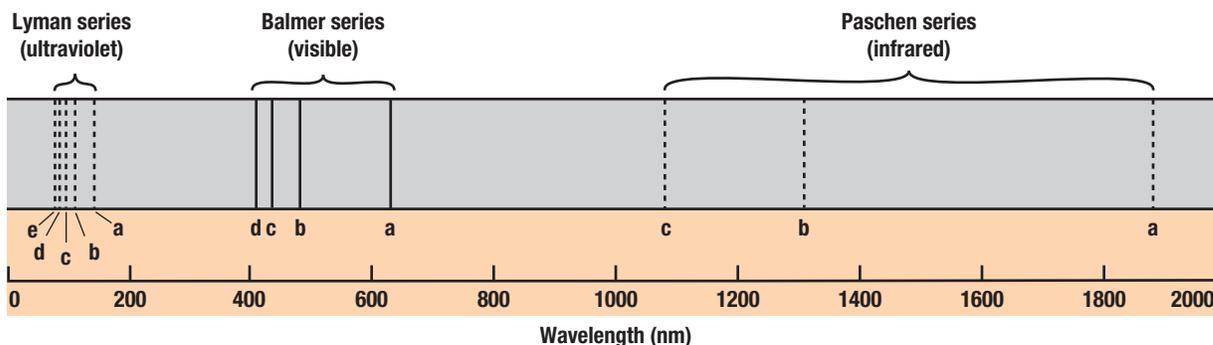

397 nm 410 nm 434 nm 486 nm 656 nm

Slits   Prism

Current is passed through a glass tube containing hydrogen at low pressure. The line at 397 nm is in the ultraviolet and is not visible to the human eye.

FIGURE 1.6

Explaining Energy-Levels A series of specific wavelengths of emitted light makes up hydrogen's emission-line spectrum. The letters below the lines label hydrogen's various energy-level transitions. Niels Bohr's model of the hydrogen atom provided an explanation for these transitions.

Lyman series (ultraviolet)   Balmer series (visible)   Paschen series (infrared)

e a   d c b   a   c   b   a
d c b

0   200   400   600   800   1000   1200   1400   1600   1800   2000

Wavelength (nm)

When investigators passed electric current through a vacuum tube containing hydrogen gas at low pressure, they observed the emission of a characteristic pinkish glow. **When a narrow beam of the emitted light was passed through a prism, it was separated into four specific colors of the visible spectrum. The four bands of light were part of what is known as hydrogen's** emission-line spectrum. The production of hydrogen's emission-line spectrum is illustrated in **Figure 1.5** (on the previous page). Additional series of lines were discovered in the ultraviolet and infrared regions of hydrogen's emission-line spectrum. The wavelengths of some of the spectral series are shown in **Figure 1.6**. They are known as the Lyman, Balmer, and Paschen series, after their discoverers.

Classical theory predicted that the hydrogen atoms would be excited by whatever amount of energy was added to them. **Scientists had thus expected to observe the emission of a continuous range of frequencies of electromagnetic radiation, that is, a** continuous spectrum. Why had the hydrogen atoms given off only specific frequencies of light? Attempts to explain this observation led to an entirely new atomic theory called *quantum theory*.

Whenever an excited hydrogen atom falls to its ground state or to a lower-energy excited state, it emits a photon of radiation. The energy of this photon ($E_{photon} = h\nu$) is equal to the difference in energy between the atom's initial state and its final state, as illustrated in **Figure 1.7**. The fact that hydrogen atoms emit only specific frequencies of light indicated that the energy differences between the atoms' energy states were fixed. This suggested that the electron of a hydrogen atom exists only in very specific energy states.

In the late nineteenth century, a mathematical formula that related the various wavelengths of hydrogen's emission-line spectrum was discovered. The challenge facing scientists was to provide a model of the hydrogen atom that accounted for this relationship.

FIGURE 1.7

Mathematical Model of Energy States When an excited atom with energy E_2 falls back to energy E_1, it releases a photon that has energy $E_2 - E_1 = E_{photon} = h\nu$.
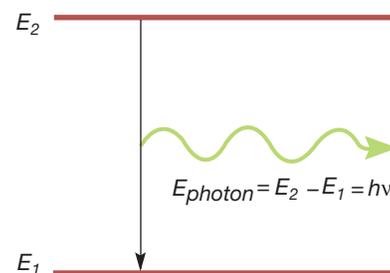
E_2

E_1

$E_{photon} = E_2 - E_1 = h\nu$

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Bohr's model of the hydrogen atom explained electron transition states.

The puzzle of the hydrogen-atom spectrum was solved in 1913 by the Danish physicist Niels Bohr. He proposed a hydrogen-atom model that linked the atom's electron to photon emission. According to the model, the electron can circle the nucleus only in allowed paths, or *orbits*. When the electron is in one of these orbits, the atom has a definite, fixed energy. The electron—and therefore the hydrogen atom—is in its lowest energy state when it is in the orbit closest to the nucleus. This orbit is separated from the nucleus by a large empty space where the electron cannot exist. The energy of the electron is higher when the electron is in orbits that are successively farther from the nucleus.

The electron orbits, or atomic energy levels, in Bohr's model can be compared to the rungs of a ladder. When you are standing on a ladder, your feet are on one rung or another. The amount of potential energy that you possess corresponds to standing on the first rung, the second rung, and so forth. Your energy cannot correspond to standing between two rungs because you cannot stand in midair. In the same way, an electron can be in one orbit or another, but not in between.

How does Bohr's model of the hydrogen atom explain the observed spectral lines? While in a given orbit, the electron is neither gaining nor losing energy. It can, however, move to a higher-energy orbit by gaining an amount of energy equal to the difference in energy between the higher-energy orbit and the initial lower-energy orbit. When a hydrogen atom is in an excited state, its electron is in one of the higher-energy orbits. When the electron falls to a lower energy level, a photon is emitted, and the process is called *emission*. The photon's energy is equal to the energy difference between the initial higher energy level and the final lower energy level. Energy must be added to an atom in order to move an electron from a lower energy level to a higher energy level. This process is called *absorption*. Absorption and emission of radiation in Bohr's model of the hydrogen atom are illustrated in **Figure 1.8**. The energy of each absorbed or emitted photon corresponds to a particular frequency of emitted radiation, $E_{photon} = h\nu$.

Based on the different wavelengths of the hydrogen emission-line spectrum, Bohr calculated the allowed energy levels for the hydrogen atom. He then related the possible energy-level changes to the lines in the hydrogen emission-line spectrum. The five lines in the Lyman series, for example, were shown to be the result of electrons dropping from energy levels E_6, E_5, E_4, E_3, and E_2 to the ground-state energy level E_1.

FIGURE 1.8

Absorption and Emission **(a)** Absorption and **(b)** emission of a photon by a hydrogen atom according to Bohr's model. The frequencies of light that can be absorbed and emitted are restricted because the electron can only be in orbits corresponding to the energies E_1, E_2, E_3, and so forth.

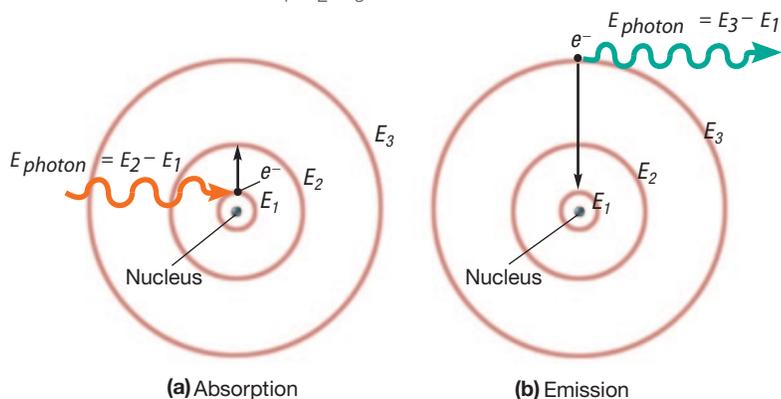
(a) Absorption

(b) Emission

FIGURE 1.9

Electron Energy Transitions This energy-state diagram for a hydrogen atom shows some of the energy transitions for the Lyman, Balmer, and Paschen spectral series. Bohr's model of the atom accounted mathematically for the energy of each of the transitions shown.

✔ **CRITICAL THINKING**

Explain Why might spectra of atoms with more than one electron be difficult to explain using Bohr's model?

E_∞
E_6
E_5
E_4
E_3
a b c
Paschen series
E_2
a b c d
Balmer series
E_1
a b c d e
Lyman series

Energy
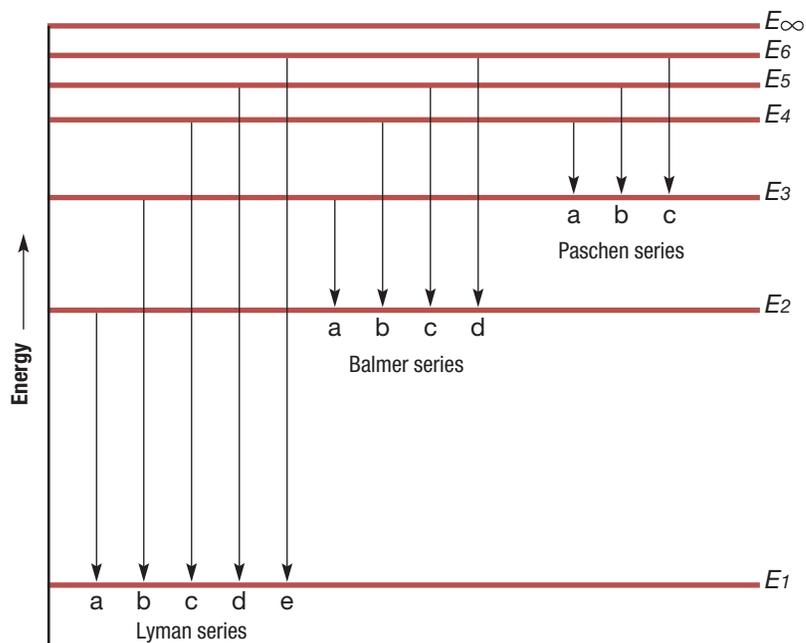
Bohr's calculated values agreed with the experimentally observed values for the lines in each series. The origins of three of the series of lines in hydrogen's emission-line spectrum are shown in **Figure 1.9.**

Bohr's model of the hydrogen atom explained observed spectral lines so well that many scientists concluded that the model could be applied to all atoms. It was soon recognized, however, that Bohr's approach did not explain the spectra of atoms with more than one electron. Nor did Bohr's theory explain the chemical behavior of atoms.

✔ ## SECTION 1 FORMATIVE ASSESSMENT

▶ **Reviewing Main Ideas**

1. What was the major shortcoming of Rutherford's model of the atom?

2. Write and label the equation that relates the speed, wavelength, and frequency of electromagnetic radiation.

3. Define the following:
   a. electromagnetic radiation
   b. wavelength
   c. frequency
   d. quantum
   e. photon

4. What is meant by the dual wave-particle nature of light?

5. Describe Bohr's model of the hydrogen atom.

✔ **Critical Thinking**

6. INTERPRETING GRAPHICS Use the diagram in **Figure 1.9** to answer the following:

   a. Characterize each of the following as absorption or emission: an electron moves from E_2 to E_1; an electron moves from E_1 to E_3; and an electron moves from E_6 to E_3.

   b. Which energy-level change above emits or absorbs the highest energy? the lowest energy?

The Quantum Model of the Atom

Key Terms

Heisenberg uncertainty principle	quantum number	magnetic quantum number
quantum theory	principal quantum number	spin quantum number
orbital	angular momentum quantum number	

To the scientists of the early twentieth century, Bohr's model of the hydrogen atom contradicted common sense. Why did hydrogen's electron exist around the nucleus only in certain allowed orbits with definite energies? Why couldn't the electron exist in a limitless number of orbits with slightly different energies? To explain why atomic energy states are quantized, scientists had to change the way they viewed the nature of the electron.

▷ **MAIN IDEA**

Electrons have wavelike properties.

The investigations into the photoelectric effect and hydrogen's emission-line spectrum revealed that light could behave as both a wave and a particle. Could electrons have a dual wave-particle nature as well? In 1924, the French scientist Louis de Broglie asked himself this very question. And the answer that he proposed led to a revolution in our basic understanding of matter.

De Broglie pointed out that in many ways the behavior of electrons in Bohr's quantized orbits was similar to the known behavior of waves. For example, scientists at the time knew that any wave confined to a space can have only certain frequencies. De Broglie suggested that electrons be considered waves confined to the space around an atomic nucleus. It followed that the electron waves could exist only at specific frequencies. And according to the relationship $E = h\nu$, these frequencies corresponded to specific energies—the quantized energies of Bohr's orbits.

Other aspects of de Broglie's hypothesis that electrons have wave-like properties were soon confirmed by experiments. Investigators demonstrated that electrons, like light waves, can be bent, or diffracted. *Diffraction* refers to the bending of a wave as it passes by the edge of an object or through a small opening. Diffraction experiments and other investigations also showed that electron beams, like waves, can interfere with each other. *Interference* occurs when waves overlap (see the Quick Lab in this section). This overlapping results in a reduction of energy in some areas and an increase of energy in others. The effects of diffraction and interference can be seen in **Figure 2.1,** on the next page.

FIGURE 2.1

Interference Patterns Bright areas correspond to areas of increased energy, while dark areas indicate areas of decreased energy.
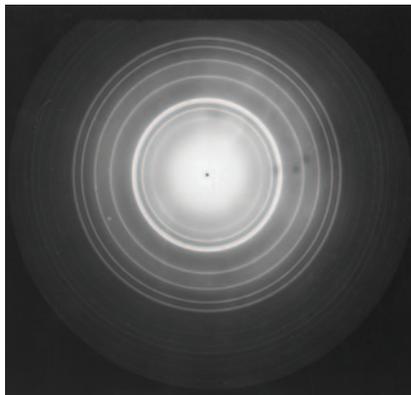
(a) Diffraction pattern produced by a beam of electrons passed through a substance.
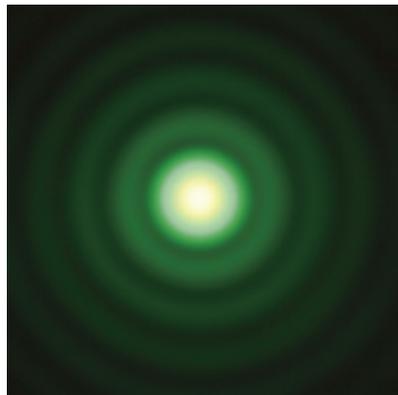
(b) Diffraction pattern for a beam of visible light passed through a tiny aperture.

✓ **CRITICAL THINKING**

Apply How did evidence of diffraction and interference indicate that electrons behave like waves of light?

▶ **MAIN IDEA**

The velocity and position of an electron cannot be measured simultaneously.

The idea of electrons having a dual wave-particle nature troubled scientists. If electrons are both particles and waves, then where are they in the atom? To answer this question, we must consider a proposal first made in 1927 by the German theoretical physicist Werner Heisenberg.

Heisenberg's idea involved the detection of electrons. Electrons are detected by their interaction with photons. Because photons have about the same energy as electrons, any attempt to locate a specific electron with a photon knocks the electron off its course. As a result, there is always a basic uncertainty in trying to locate an electron (or any other particle). **The Heisenberg uncertainty principle states that it is impossible to determine simultaneously both the position and velocity of an electron or any other particle.** This was a difficult idea for scientists to accept at the time, but it is now fundamental to our understanding of light and matter.

▶ **MAIN IDEA**

Orbitals indicate probable electron locations.

In 1926, the Austrian physicist Erwin Schrödinger used the hypothesis that electrons have a dual wave-particle nature to develop an equation that treated electrons in atoms as waves. Unlike Bohr's theory, which assumed quantization as a fact, quantization of electron energies was a natural outcome of Schrödinger's equation. Only waves of specific energies, and therefore frequencies, provided solutions to the equation. Along with the uncertainty principle, the Schrödinger wave equation laid the foundation for modern quantum theory. **Quantum theory describes mathematically the wave properties of electrons and other very small particles.**

QUESTION

Does light show the wave property of interference when a beam of light is projected through a pinhole onto a screen?

PROCEDURE

Record all your observations.

1. To make the pinhole screen, cut a 20 cm × 20 cm square from a manila folder. In the center of the square, cut a 2 cm square hole. Cut a 7 cm × 7 cm square of aluminum foil. Using a thumbtack, make a pinhole in the center of the foil square. Tape the aluminum foil over the 2 cm square hole, making sure the pinhole is centered as shown in the diagram.

2. Use white poster board to make a projection screen 35 cm × 35 cm.

3. In a dark room, center the light beam from a flashlight on the pinhole. Hold the flashlight about 1 cm from the pinhole. The pinhole screen should be about 50 cm from the projection screen, as shown in the diagram. Adjust the distance to form a sharp image on the projection screen.

DISCUSSION

1. Did you observe interference patterns on the screen?

2. As a result of your observations, what do you conclude about the nature of light?

MATERIALS

- scissors
- manila folders
- thumbtack
- masking tape
- aluminum foil
- white poster board or cardboard
- flashlight

SAFETY

Wear safety goggles and an apron.

1 cm

Image

50 cm
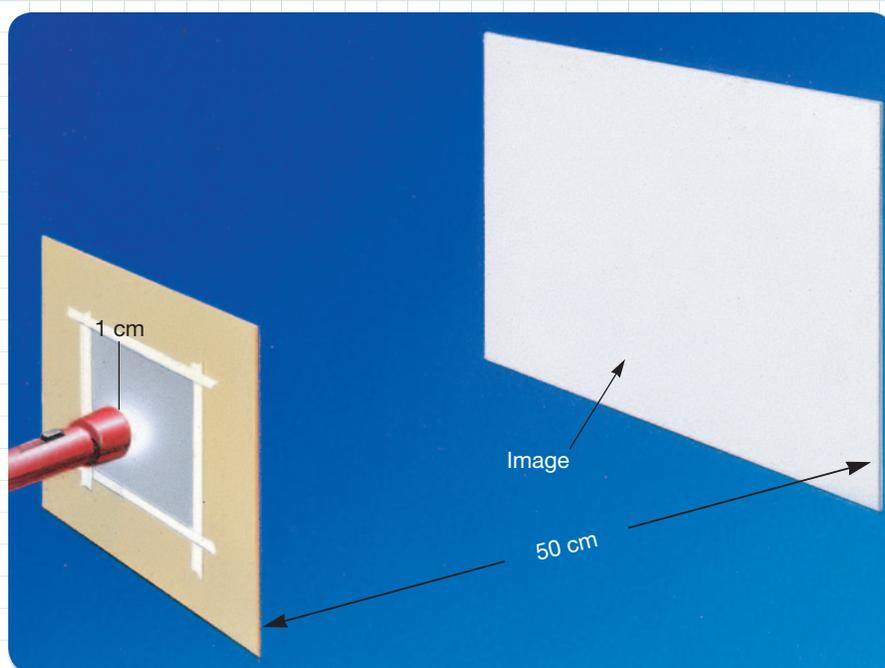
Solutions to the Schrödinger wave equation are known as wave functions. Based on the Heisenberg uncertainty principle, the early developers of quantum theory determined that wave functions give only the *probability* of finding an electron at a given place around the nucleus. Thus, electrons do not travel around the nucleus in neat orbits, as Bohr had postulated. Instead, they exist in certain regions called orbitals. **An orbital is a three-dimensional region around the nucleus that indicates the probable location of an electron. Figure 2.2**, on the next page, illustrates two ways of picturing one type of atomic orbital. As you will see later in this section, atomic orbitals have different shapes and sizes.

FIGURE 2.2

Atomic Orbitals Two ways of showing a simple atomic orbital are presented.
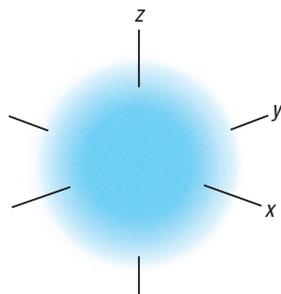
(a) In this atomic orbital model, the probability of finding the electron is proportional to the density of the cloud.
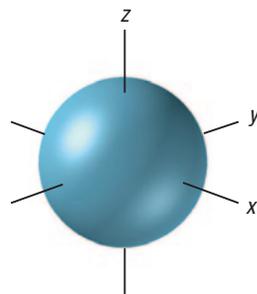
(b) An orbital can also be depicted as a surface within which the electron can be found a certain percentage of the time, conventionally 90%.

▶ **MAIN IDEA**

Quantum numbers describe atomic orbitals.

According to the Schrödinger equation, electrons in atomic orbitals also have quantized energies. Unlike in the Bohr atomic model, therefore, an electron's energy level is not the only characteristic of an orbital.

To describe orbitals accurately, scientists use quantum numbers. **Quantum numbers** specify the properties of atomic orbitals and the properties of electrons in orbitals. The first three quantum numbers result from solutions to the Schrödinger equation. They indicate the main energy level, the shape, and the orientation of an orbital. The fourth, the spin quantum number, describes a fundamental state of the electron that occupies the orbital.

Principal Quantum Number

The **principal quantum number,** symbolized by *n,* indicates the main energy level occupied by the electron. Values of n are positive integers only—1, 2, 3, and so on. As n increases, the electron's energy and its average distance from the nucleus increase, as shown in **Figure 2.3**. For example, an electron for which $n = 1$ occupies the first, or lowest, main energy level and is located closest to the nucleus. An electron for which $n = 2$ occupies the second energy level and is farther from the nucleus.

More than one electron can have the same n value. These electrons are said to be in the same electron *shell.* The total number of orbitals in a given shell, or main energy level, is equal to n^2.

FIGURE 2.3

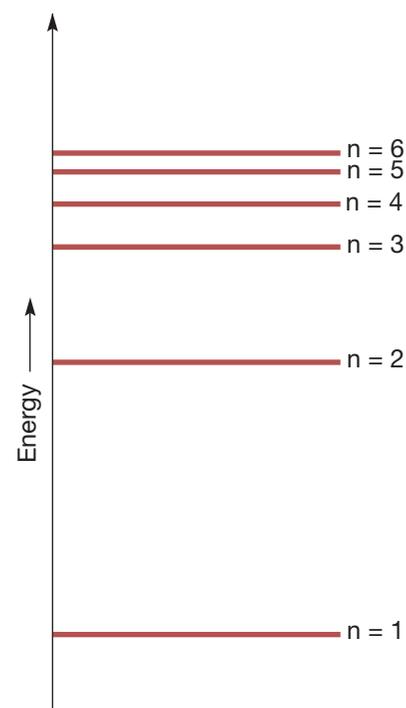**The Principle Quantum Number** The main energy levels of an atom are represented by the principal quantum number, *n.*

ORBITAL LETTER DESIGNATIONS

ℓ	Letter
0	*s*
1	*p*
2	*d*
3	*f*

Angular Momentum Quantum Number

Except at the first main energy level, orbitals of different shapes exist for a given value of *n*. These are known as *sublevels*. **The angular momentum quantum number, symbolized by ℓ, indicates the shape of the orbital.** For a specific main energy level, the number of orbital shapes possible is equal to *n*. The values of ℓ allowed are zero and all positive integers less than or equal to $n - 1$. For example, orbitals for which $n = 2$ can have one of two shapes corresponding to $\ell = 0$ and $\ell = 1$. Depending on its value of ℓ, an orbital is assigned a letter, such as *s, p,* and *d*. Orbital letter designations for values of ℓ are given in **Figure 2.4.**

As shown in **Figure 2.5a** on the next page, *s* orbitals are spherical, *p* orbitals have dumbbell shapes, and *d* orbitals are more complex. (The *f* orbital shapes are even more complex.) In the first energy level, $n = 1$, there is only one sublevel possible—an *s* orbital. As mentioned, the second energy level, $n = 2$, has two sublevels—the *s* and *p* orbitals. The third energy level, $n = 3$, has three sublevels—the *s, p,* and *d* orbitals. The fourth energy level, $n = 4$, has four sublevels—the *s, p, d,* and *f* orbitals. In an *n*th main energy level, there are *n* sublevels.

Each atomic orbital is designated by the principal quantum number followed by the letter of the sublevel. For example, the 1*s* sublevel is the *s* orbital in the first main energy level, while the 2*p* sublevel is the set of three *p* orbitals in the second main energy level. On the other hand, a 4*d* orbital is part of the *d* sublevel in the fourth main energy level. How would you designate the *p* sublevel in the third main energy level? How many other sublevels are in the third main energy level with this one?

Magnetic Quantum Number

Atomic orbitals can have the same shape but different orientations around the nucleus. **The magnetic quantum number, symbolized by *m*, indicates the orientation of an orbital around the nucleus.** Values of *m* are intergers, including zero, from $-\ell$ to $+\ell$. Because an *s* orbital is spherical and is centered around the nucleus, it has only one possible orientation. This orientation corresponds to a magnetic quantum number of $m = 0$. There is therefore only one *s* orbital in each *s* sublevel.

As shown in **Figure 2.5b** on the next page, the lobes of a *p* orbital extend along the *x*-, *y*-, or *z*-axis of a three-dimensional coordinate system. There are therefore three *p* orbitals in each *p* sublevel, designated as p_x, p_y, and p_z orbitals. The three *p* orbitals occupy different regions of space, and those regions are related to values of $m = -1$, $m = 0$, and $m = +1$.

There are five different *d* orbitals in each *d* sublevel, as shown in **Figure 2.5c** on the next page. For the *d* orbital, the five different orientations, including one with a different shape, correspond to values of $m = -2$, $m = -1$, $m = 0$, $m = +1$, and $m = +2$. There are seven different *f* orbitals in each *f* sublevel. (The *f* orbitals are not shown in **Figure 2.5.**)

FIGURE 2.5

a. Orbital Shapes

The orbitals *s*, *p*, and *d* have different shapes. Each of the orbitals shown occupies a different region of space around the nucleus.

s orbital
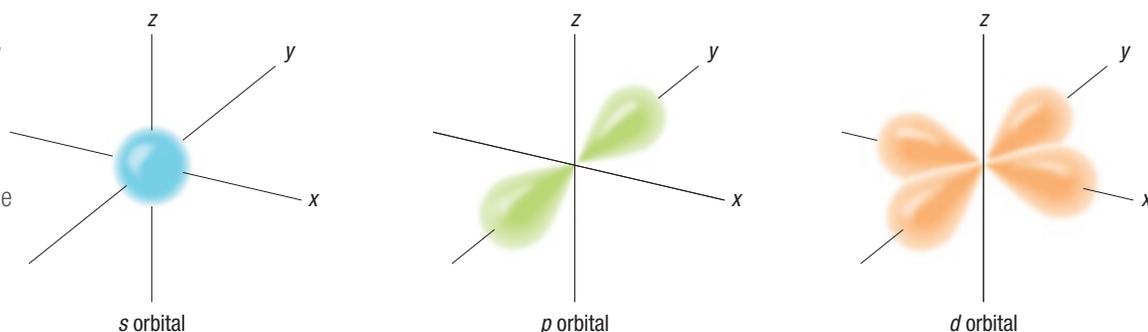
p orbital

d orbital

b. Three Orientations of *p* Orbitals

The subscripts *x*, *y*, and *z* indicate the three orientations of *p* orbitals. The intersection of the *x*-, *y*-, and *z*-axes indicates the location of the center of the nucleus.

p_x orbital

p_y orbital
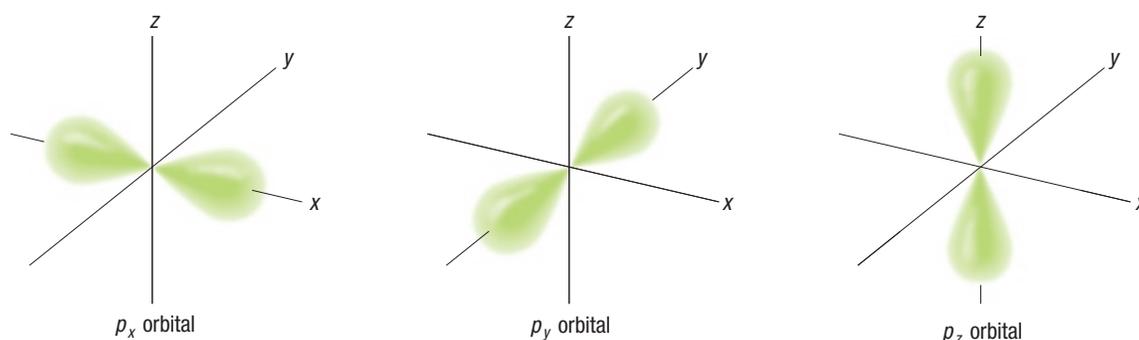
p_z orbital

c. Five Orientations of *d* Orbitals

Four of the *d* orbitals have the same shape but different orientations. The fifth has a different shape and a different orientation. Each orbital occupies a different region of space.

$d_{x^2-y^2}$ orbital

d_{xy} orbital

d_{yz} orbital

d_{xz} orbital

d_{z^2} orbital
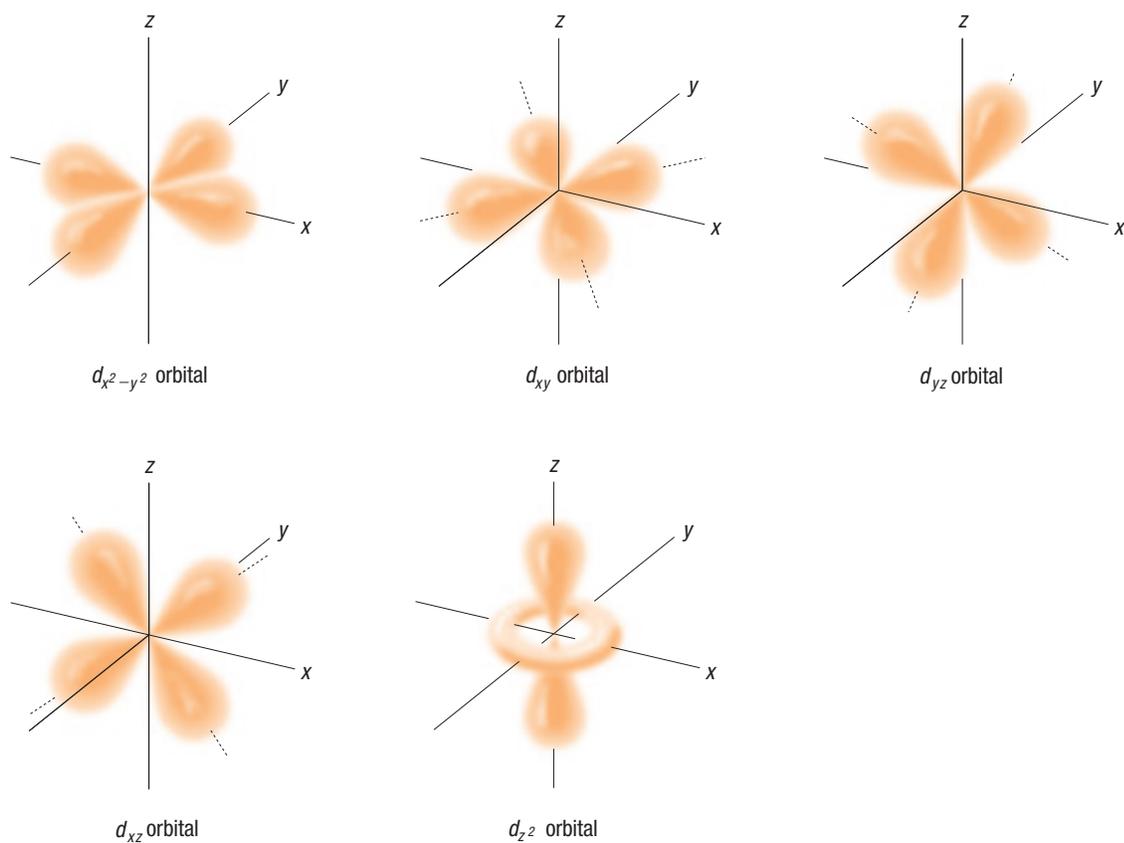
FIGURE 2.6

QUANTUM NUMBER RELATIONSHIPS IN ATOMIC STRUCTURE

Principal quantum number: main energy level (n)	Sublevel in main energy level (n sublevels)	Number of orbitals per sublevel	Number of orbitals per main energy level (n^2)	Number of electrons per sublevel	Number of electrons per main energy level ($2n^2$)
1	s	1	1	2	2
2	s p	1 3	4	2 6	8
3	s p d	1 3 5	9	2 6 10	18
4	s p d f	1 3 5 7	16	2 6 10 14	32

As you can see in **Figure 2.6**, the total number of orbitals in a main energy level increases with the value of n. In fact, the number of orbitals at each main energy level equals the square of the principal quantum number, n^2.

Spin Quantum Number

An electron in an orbital behaves in some ways like Earth spinning on an axis. The electron exists in one of two possible spin states, which creates a magnetic field. To account for the magnetic properties of the electron, theoreticians of the early twentieth century created the spin quantum number. **The spin quantum number has only two possible values— ($+1/2$, $-1/2$)—which indicate the two fundamental spin states of an electron in an orbital.** A single orbital can hold a maximum of two electrons, but the two electrons must have opposite spin states.

CHECK FOR UNDERSTANDING

Analyze Why are quantum numbers used to describe electrons in atoms?

✔ SECTION 2 FORMATIVE ASSESSMENT

▶ Reviewing Main Ideas

1. Define the following:
   a. main energy levels
   b. quantum numbers

2. a. List the four quantum numbers.
   b. What general information about atomic orbitals is provided by the quantum numbers?

3. Describe briefly what specific information is given by each of the four quantum numbers.

✔ Critical Thinking

4. INFERRING RELATIONSHIPS What are the possible values of the magnetic quantum number m for f orbitals? What is the maximum number of electrons that can exist in $4f$ orbitals?

Electron Configurations

Key Terms

electron configuration
Aufbau principle

Pauli exclusion principle
Hund's rule

noble gas
noble-gas configuration

The quantum model of the atom improves on the Bohr model because it describes the arrangements of electrons in atoms other than hydrogen. **The arrangement of electrons in an atom is known as the atom's electron configuration.** Because atoms of different elements have different numbers of electrons, a unique electron configuration exists for the atoms of each element. Like all systems in nature, electrons in atoms tend to assume arrangements that have the lowest possible energies. The lowest-energy arrangement of the electrons for each element is called the element's *ground-state electron configuration*. A few simple rules, combined with the quantum number relationships discussed in Section 2, allow us to determine these ground-state electron configurations.

▶ **MAIN IDEA**
Electrons fill in the lowest-energy orbitals first.

To build up electron configurations for the ground state of any particular atom, first the energy levels of the orbitals are determined. Then electrons are added to the orbitals, one by one, according to three basic rules. (Remember that real atoms are not built up by adding protons and electrons one at a time.)

The first rule shows the order in which electrons occupy orbitals. **According to the Aufbau principle, an electron occupies the lowest-energy orbital that can receive it.** **Figure 3.1** shows the atomic orbitals in order of increasing energy. The orbital with the lowest energy is the $1s$ orbital. In a ground-state hydrogen atom, the electron is in this orbital. The $2s$ orbital is the next highest in energy, then the $2p$ orbitals. Beginning with the third main energy level, $n = 3$, the energies of the sublevels in different main energy levels begin to overlap.

Note in the figure, for example, that the $4s$ sublevel is lower in energy than the $3d$ sublevel. Therefore, the $4s$ orbital is filled before any electrons enter the $3d$ orbitals. (Less energy is required for two electrons to pair up in the $4s$ orbital than for those two electrons to occupy a $3d$ orbital.) Once the $3d$ orbitals are fully occupied, which sublevel will be occupied next?

FIGURE 3.1

Atomic Sublevels The order of increasing energy for atomic sublevels is shown on the vertical axis. Each individual box represents an orbital.
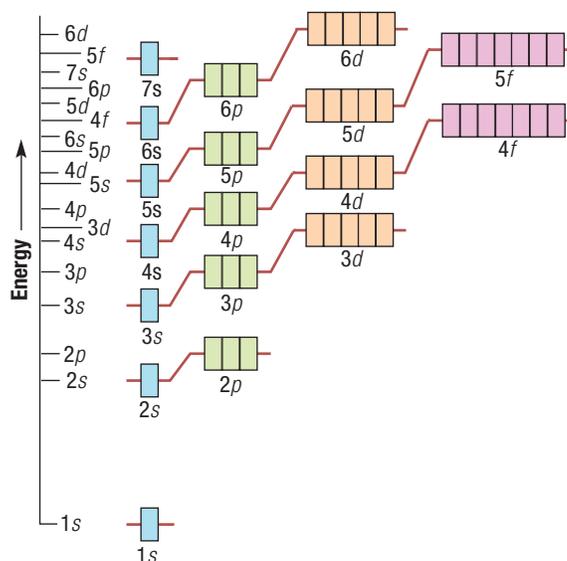
FIGURE 3.2

Pauli Exclusion Principle

According to the Pauli exclusion principle, an orbital can hold two electrons of opposite spin states. In this electron configuration of a helium atom, each arrow represents one of the atom's two electrons. The direction of the arrow indicates the electron's spin state.
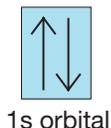
1s orbital

The second rule reflects the importance of the spin quantum number. **According to the Pauli exclusion principle, no two electrons in the same atom can have the same set of four quantum numbers.** The principal, angular momentum, and magnetic quantum numbers specify the energy, shape, and orientation of an orbital. The two values of the spin quantum number reflect the fact that for two electrons to occupy the same orbital, they must have opposite spin states (see **Figure 3.2**).

The third rule requires placing as many unpaired electrons as possible in separate orbitals in the same sublevel. In this way, electron-electron repulsion is minimized so that the electron arrangements have the lowest energy possible. **According to Hund's rule, orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons in singly occupied orbitals must have the same spin state.** Applying this rule shows, for example, that one electron will enter each of the three p orbitals in a main energy level before a second electron enters any of them. This is illustrated in **Figure 3.3**. What is the maximum number of unpaired electrons in a d sublevel?

FIGURE 3.3

Hund's Rule   The figure shows how **(a)** two, **(b)** three, and **(c)** four electrons fill the p sublevel of a given main energy level according to Hund's rule.

(a)   (b)   (c)
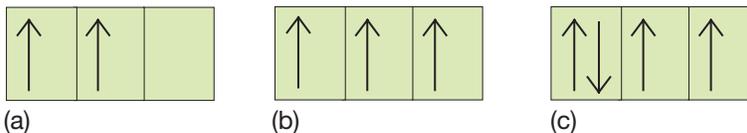
▶ **MAIN IDEA**

There are three ways to indicate electron configuration.

Three methods, or notations, are used to indicate electron configurations. Two of these notations will be discussed for the first-period elements, hydrogen and helium, on the following page. The third notation is used mostly with elements of the third period and higher and will be discussed later in this section.

In a ground-state hydrogen atom, the single electron is in the lowest-energy orbital, the 1s orbital. The electron can be in either one of its two spin states. Helium has two electrons, which are paired in the 1s orbital.

Orbital Notation

In orbital notation, an unoccupied orbital is represented by a line, ⎯ , with the orbital's name written underneath the line. An orbital containing one electron is represented as ↑. An orbital containing two electrons is represented as ↑↓ , showing the electrons paired and with opposite spin states. The lines are labeled with the principal quantum number and sublevel letter. For example, the orbital notations for hydrogen and helium are written as follows:

$$H \frac{\uparrow}{1s} \qquad He \frac{\uparrow\downarrow}{1s}$$

Electron-Configuration Notation

Electron-configuration notation eliminates the lines and arrows of orbital notation. Instead, the number of electrons in a sublevel is shown by adding a superscript to the sublevel designation. The hydrogen configuration is represented by $1s^1$. The superscript indicates that one electron is present in hydrogen's $1s$ orbital. The helium configuration is represented by $1s^2$. Here the superscript indicates that there are two electrons in helium's $1s$ orbital.

Electron Configurations

Sample Problem A The electron configuration of boron is $1s^2 2s^2 2p^1$. How many electrons are present in an atom of boron? What is the atomic number for boron? Write the orbital notation for boron.

● SOLVE

The number of electrons in a boron atom is equal to the sum of the superscripts in its electron-configuration notation: $2 + 2 + 1 = 5$ electrons. The number of protons equals the number of electrons in a neutral atom. So we know that boron has 5 protons and thus has an atomic number of 5. To write the orbital notation, first draw the lines representing orbitals.

$$\underset{1s}{\quad} \underset{2s}{\quad} \quad \underset{2p}{\underbrace{\quad\ \quad\ \quad}}$$
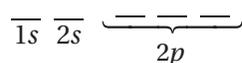
Next, add arrows showing the electron locations. The first two electrons occupy $n = 1$ energy level and fill the $1s$ orbital.

$$\underset{1s}{\frac{\uparrow\downarrow}{\quad}} \underset{2s}{\quad} \quad \underset{2p}{\underbrace{\quad\ \quad\ \quad}}$$
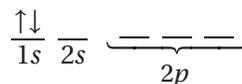
The next three electrons occupy the $n = 2$ main energy level. Two of these occupy the lower-energy $2s$ orbital. The third occupies a higher-energy p orbital.

$$\underset{1s}{\frac{\uparrow\downarrow}{\quad}} \underset{2s}{\frac{\uparrow\downarrow}{\quad}} \quad \underset{2p}{\underbrace{\frac{\uparrow}{\quad}\ \quad\ \quad}}$$
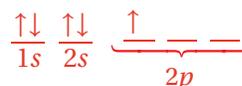
Practice Answers in Appendix E

1. The electron configuration of nitrogen is $1s^2 2s^2 2p^3$. How many electrons are in a nitrogen atom? What is the atomic number of nitrogen? Write the orbital notation for nitrogen.
2. The electron configuration of fluorine is $1s^2 2s^2 2p^5$. What is the atomic number of fluorine? How many of its p orbitals are filled?

The Noble Decade

By the late nineteenth century, the science of chemistry had begun to be organized. In 1860, the First International Congress of Chemistry established the field's first standards. And Dmitri Mendeleev's periodic table of elements gave chemists across the globe a systematic understanding of matter's building blocks. But many important findings—including the discovery of a family of rare, unreactive gases that were unlike any substances known at the time—were yet to come.

Cross-Disciplinary Correspondence

In 1888, the British physicist Lord Rayleigh encountered a small but significant discrepancy in the results of one of his experiments. In an effort to redetermine the atomic mass of nitrogen, he measured the densities of several samples of nitrogen gas. Each sample had been prepared by a different method. All samples that had been isolated from chemical reactions exhibited similar densities. But the densities were about one-tenth of a percent less dense than the nitrogen isolated from air, which at the time was believed to be a mixture of nitrogen, oxygen, water vapor, and carbon dioxide.

Rayleigh was at a loss to explain his discovery. Finally, in 1892, he published a letter in *Nature* magazine to appeal to his colleagues for an explanation. A month later, he received a reply from a Scottish chemist named William Ramsay. Ramsay related that he too had been stumped by the density difference between chemical and atmospheric nitrogen. Rayleigh decided to report his findings to the Royal Society.

A Chemist's Approach

With Rayleigh's permission, Ramsay attempted to remove all known components from a sample of air and to analyze what, if anything, remained. Having removed water vapor, carbon dioxide, and oxygen from the air, Ramsay repeatedly passed the sample over hot magnesium. The nitrogen reacted with the magnesium to form solid magnesium nitride. As a result, all of the then-known components of air were removed. What remained was a minuscule portion of a mysterious gas.

Ramsay tried to cause the gas to react with chemically active substances, such as hydrogen, sodium, and caustic soda, but the gas remained unaltered. He decided to name this new atmospheric component *argon* (Greek for "inert" or "idle").

Periodic Problems

Rayleigh and Ramsay were sure that they had discovered a new element, which created a problem. Their calculations indicated that argon had an atomic mass of about 40. However, as it appeared in 1894, the periodic table had no space for such an element. The elements with atomic masses closest to that of argon were chlorine and potassium. Unfortunately, the chemical properties of the families of each of these elements were completely dissimilar to those of the strange gas.

SEPTEMBER 29, 1892   *NATURE*

LETTERS TO THE EDITOR.

[*The Editor does not hold himself responsible for opinions ex-pressed by his correspondents. Neither can he undertake to return, or to correspond with the writers of, rejected manuscripts intended for this or any other part of* NATURE. *No notice is taken of anonymous communications.*]

Density of Nitrogen.

I AM much puzzled by some recent results as to the density of *nitrogen*, and shall be obliged if any of your chemical readers can offer suggestions as to the cause. According to two methods of preparation I obtain quite distinct values. The relative difference, amounting to about $\frac{1}{1000}$ part, is small in itself; but it lies entirely outside the errors of experiment, and can only be attributed to a variation in the character of the gas...

Is it possible that the difference is independent of impurity, the nitrogen itself being to some extent in a different (dissociated) state ?...

RAYLEIGH.

Terling Place, Witham, September 24.

This excerpt from Lord Rayleigh's letter was originally published in **Nature** *magazine in 1892.*

Groups	III	IV	V	VI	VII	VIII	I	II	III	IV	V	VI	VII	0	I	II
Periods	b	b	b	b	b	b	b	b	a	a	a	a	a		a	a
1															H	He
2															Li	Be
3									B	C	N	O	F	Ne	Na	Mg
4									Al	Si	P	S	Cl	Ar	K	Ca
5	Sc	Ti	V	Cr	Mn	Fe Co Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	Rb	Sr
6	Y	Zr	Nb	Mo	Tc	Ru Rh Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	Cs	Ba
7	La	Hf	Ta	W	Re	Os Ir Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	Fr	Ra
8	Ac															
				Transition elements							Main-group elements					
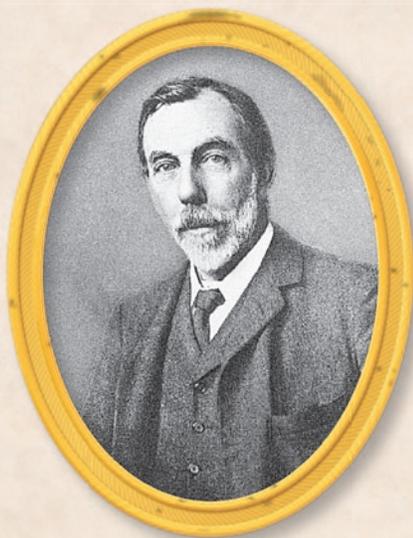
In 1893, Scottish chemist William Ramsay isolated a previously unknown component of the atmosphere.

This version of the periodic table shows how it looked after the discovery of the noble gases. The placement of the Group 1 and 2 elements at the far right of the table clearly illustrates how the noble gases fit in between the chlorine family and the potassium family of elements. The 0 above the noble-gas family indicates the zero valency of the gases.

Ramsay contemplated argon's lack of reactivity. He knew that Mendeleev had created the periodic table on the basis of valence, or the number of atomic partners an element bonds with in forming a compound. Because Ramsay could not cause argon to form any compounds, he assigned it a valence of zero. And because the valence of the elements in the families of both chlorine and potassium was one, perhaps argon fit in between them.

Ramsay's insight that argon merited a new spot between the halogen family and the alkali metal family on the periodic table was correct. And as Ramsay would soon confirm, his newly discovered gas was indeed one of a previously unknown family of elements.

New Neighbors

In 1895, Ramsay isolated a light, inert gas from a mineral called *cleveite*. Physical analysis revealed that the gas was the same as one that had been identified in the sun in 1868—helium. Helium was the second zero-valent element found on Earth, and its discovery made chemists aware that the periodic table had been missing an entire column of elements.

Over the next three years, Ramsay and his assistant, Morris Travers, identified three more inert gases present in the atmosphere: neon (Greek for "new"), krypton ("hidden"), and xenon ("stranger"). Finally in 1900, German chemist Friedrich Ernst Dorn discovered radon, the last of the new family of elements known today as the *noble gases*. For his discovery, Ramsay received the Nobel Prize in 1904.

Questions

1. What evidence led Ramsay to report that the mysterious gas was inert?

2. What property of argon caused Ramsay to propose a new column in the periodic table?

MAIN IDEA

No electrons can occupy a higher-energy sublevel until the sublevel below it is filled.

In the first-period elements, hydrogen and helium, electrons occupy the orbital of the first main energy level. **Figure 3.4** provides a pattern to help you remember the order in which orbitals are filled according to the Aufbau principle. The ground-state configurations in **Figure 3.5** illustrate how the Aufbau principle, the Pauli exclusion principle, and Hund's rule are applied to atoms of elements in the second period.

According to the Aufbau principle, after the 1s orbital is filled, the next electron occupies the s sublevel in the second main energy level. Thus, lithium, Li, has a configuration of $1s^2 2s^1$. The electron occupying the 2s level of a lithium atom is in the atom's highest, or outermost, occupied level. The *highest-occupied energy level* is the electron-containing main energy level with the highest principal quantum number. The two electrons in the 1s sublevel of lithium are no longer in the outermost main energy level. They have become *inner-shell electrons*, which are electrons that are not in the highest-occupied energy level.

The fourth electron in an atom of beryllium, Be, must complete the pair in the 2s sublevel because this sublevel is of lower energy than the 2p sublevel. With the 2s sublevel filled, the 2p sublevel, which has three vacant orbitals of equal energy, can be occupied. One of the three p orbitals is occupied by a single electron in an atom of boron, B. Two of the three p orbitals are occupied by unpaired electrons in an atom of carbon, C. And all three p orbitals are occupied by unpaired electrons in an atom of nitrogen, N. Hund's rule applies here, as is shown in the orbital notations in **Figure 3.5**.

FIGURE 3.4

The Aufbau Principle Follow the diagonal arrows from the bottom of one column to the top of the next to get the order in which atomic orbitals are filled according to the Aufbau principle.
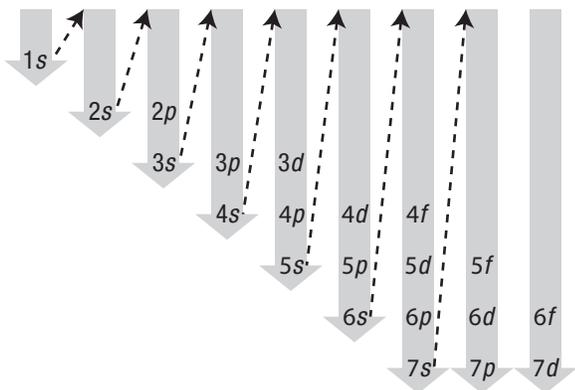
✔️ **CRITICAL THINKING**

Interpret Review the statement of the Aufbau principle given at the beginning of this section, and then study Figure 3.4 carefully. Based on your understanding of the two, explain in your own words why electrons fill atomic orbitals in this seemingly unusual order.

According to the Aufbau principle, the next electron must pair with another electron in one of the 2p orbitals rather than enter the third main energy level. The Pauli exclusion principle allows the electron to pair with one of the electrons occupying the 2p orbitals as long as the spins of the paired electrons are opposite. Thus, atoms of oxygen, O, have the configuration $1s^2 2s^2 2p^4$. Oxygen's orbital notation is shown in **Figure 3.5**.

Two 2p orbitals are filled in fluorine, F, and all three are filled in neon, Ne. Atoms such as those of neon, which have the s and p sublevels of their highest occupied level filled with eight electrons, are said to have an *octet* of electrons. Examine the periodic table inside the back cover of the text. Notice that neon is the last element in the second period.

FIGURE 3.5

ELECTRON CONFIGURATIONS OF ATOMS OF SECOND-PERIOD ELEMENTS SHOWING TWO NOTATIONS

Name	Symbol	1s	2s	2p			Electron-configuration notation
Lithium	Li	↑↓	↑	—	—	—	$1s^2 2s^1$
Beryllium	Be	↑↓	↑↓	—	—	—	$1s^2 2s^2$
Boron	B	↑↓	↑↓	↑	—	—	$1s^2 2s^2 2p^1$
Carbon	C	↑↓	↑↓	↑	↑	—	$1s^2 2s^2 2p^2$
Nitrogen	N	↑↓	↑↓	↑	↑	↑	$1s^2 2s^2 2p^3$
Oxygen	O	↑↓	↑↓	↑↓	↑	↑	$1s^2 2s^2 2p^4$
Fluorine	F	↑↓	↑↓	↑↓	↑↓	↑	$1s^2 2s^2 2p^5$
Neon	Ne	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6$

Elements of the Third Period

After the outer octet is filled in neon, the next electron enters the s sublevel in the $n = 3$ main energy level. Thus, atoms of sodium, Na, have the configuration $1s^2 2s^2 2p^6 3s^1$. Compare the configuration of a sodium atom with that of an atom of neon in **Figure 3.5.** Notice that the first 10 electrons in a sodium atom have the same configuration as a neon atom, $1s^2 2s^2 2p^6$. In fact, the first 10 electrons in an atom of each of the third-period elements have the same configuration as neon. This similarity allows us to use a shorthand notation for the electron configurations of the third-period elements.

Noble-Gas Notation

Neon is a member of the Group 18 elements. **The Group 18 elements (helium, neon, argon, krypton, xenon, and radon) are called the noble gases.** To simplify sodium's notation, the symbol for neon, enclosed in square brackets, is used to represent the complete neon configuration: $[Ne] = 1s^2 2s^2 2p^6$. This allows us to write sodium's electron configuration as $[Ne]3s^1$, which is called sodium's *noble-gas notation*. **Figure 3.6,** on the next page, shows the electron configuration of each of the third-period elements using noble-gas notation.

The last element in the third period is argon, Ar, which is a noble gas. As in neon, the highest-occupied energy level of argon has an octet of electrons, $[Ne]3s^2 3p^6$. In fact, each noble gas other than He has an electron octet in its highest energy level. **A noble-gas configuration refers to an outer main energy level occupied, in most cases, by eight electrons.**

FIGURE 3.6

ELECTRON CONFIGURATIONS OF ATOMS OF THIRD-PERIOD ELEMENTS

Name	Symbol	Atomic number	Number of electrons in sublevels					Noble-gas notation
			1s	2s	2p	3s	3p	
Sodium	Na	11	2	2	6	1		*[Ne]$3s^1$
Magnesium	Mg	12	2	2	6	2		[Ne]$3s^2$
Aluminum	Al	13	2	2	6	2	1	[Ne]$3s^2 3p^1$
Silicon	Si	14	2	2	6	2	2	[Ne]$3s^2 3p^2$
Phosphorus	P	15	2	2	6	2	3	[Ne]$3s^2 3p^3$
Sulfur	S	16	2	2	6	2	4	[Ne]$3s^2 3p^4$
Chlorine	Cl	17	2	2	6	2	5	[Ne]$3s^2 3p^5$
Argon	Ar	18	2	2	6	2	6	[Ne]$3s^2 3p^6$

*[Ne] $= 1s^2 2s^2 2p^6$

Elements of the Fourth Period

The electron configurations of atoms in the fourth-period elements are shown in **Figure 3.7** (on the next page). The period begins by filling the $4s$ orbital, the empty orbital of lowest energy. The first element in the fourth period is potassium, K, which has the electron configuration [Ar]$4s^1$. The next element is calcium, Ca, which has the electron configuration [Ar]$4s^2$.

With the $4s$ sublevel filled, the $4p$ and $3d$ sublevels are the next available vacant orbitals. **Figure 3.1** (on the first page of this section) shows that the $3d$ sublevel is lower in energy than the $4p$ sublevel. Therefore, the five $3d$ orbitals are next to be filled. A total of 10 electrons can occupy the $3d$ orbitals. These are filled successively in the 10 elements from scandium (atomic number 21) to zinc (atomic number 30).

Scandium, Sc, has the electron configuration [Ar]$3d^1 4s^2$. Titanium, Ti, has the configuration [Ar]$3d^2 4s^2$. And vanadium, V, has the configuration [Ar]$3d^3 4s^2$. Up to this point, three electrons with the same spin have been added to three separate d orbitals, as required by Hund's rule.

Surprisingly, chromium, Cr, has the electron configuration [Ar]$3d^5 4s^1$. Not only did the added electron go into the fourth $3d$ orbital, but an electron also moved from the $4s$ orbital into the fifth $3d$ orbital, leaving the $4s$ orbital with a single electron. Chromium's electron configuration is contrary to what is expected according to the Aufbau principle. However, in reality the [Ar]$3d^5 4s^1$ configuration is of lower energy than a [Ar]$3d^4 4s^2$ configuration. For chromium, having six orbitals with unpaired electrons is a more stable arrangement than having four unpaired electrons in the $3d$ orbitals and forcing two electrons to pair up in the $4s$ orbital. On the other hand, for tungsten, W, which is in the same group as chromium, having four electrons in the $5d$ orbitals and two electrons paired in the $6s$ orbital is the most stable arrangement. There is no simple explanation for such deviations from the expected order given in **Figure 3.4**.

Manganese, Mn, has the electron configuration $[Ar]3d^54s^2$. The added electron goes to the $4s$ orbital, completely filling this orbital while leaving the $3d$ orbitals still half-filled. Beginning with the next element, electrons continue to pair in the d orbitals. Thus, iron, Fe, has the configuration $[Ar]3d^64s^2$; cobalt, Co, has the configuration $[Ar]3d^74s^2$; and nickel, Ni, has the configuration $[Ar]3d^84s^2$. Next is copper, Cu, in which an electron moves from the $4s$ orbital to pair with the electron in the fifth $3d$ orbital. The result is an electron configuration of $[Ar]3d^{10}4s^1$—the lowest-energy configuration for Cu. As with Cr, there is no simple explanation for this deviation from the expected order.

In atoms of zinc, Zn, the $4s$ sublevel is filled to give the electron configuration $[Ar]3d^{10}4s^2$. In atoms of the next six elements, electrons add one by one to the three $4p$ orbitals. According to Hund's rule, one electron is added to each of the three $4p$ orbitals before electrons are paired in any $4p$ orbital.

FIGURE 3.7

ELECTRON CONFIGURATION OF ATOMS OF ELEMENTS IN THE FOURTH PERIOD

Name	Symbol	Atomic number	Number of electrons in sublevels above $2p$					Noble-gas notation
			3s	**3p**	**3d**	**4s**	**4p**	
Potassium	K	19	2	6		1		*$[Ar]4s^1$
Calcium	Ca	20	2	6		2		$[Ar]4s^2$
Scandium	Sc	21	2	6	1	2		$[Ar]3d^14s^2$
Titanium	Ti	22	2	6	2	2		$[Ar]3d^24s^2$
Vanadium	V	23	2	6	3	2		$[Ar]3d^34s^2$
Chromium	Cr	24	2	6	5	1		$[Ar]3d^54s^1$
Manganese	Mn	25	2	6	5	2		$[Ar]3d^54s^2$
Iron	Fe	26	2	6	6	2		$[Ar]3d^64s^2$
Cobalt	Co	27	2	6	7	2		$[Ar]3d^74s^2$
Nickel	Ni	28	2	6	8	2		$[Ar]3d^84s^2$
Copper	Cu	29	2	6	10	1		$[Ar]3d^{10}4s^1$
Zinc	Zn	30	2	6	10	2		$[Ar]3d^{10}4s^2$
Gallium	Ga	31	2	6	10	2	1	$[Ar]3d^{10}4s^24p^1$
Germanium	Ge	32	2	6	10	2	2	$[Ar]3d^{10}4s^24p^2$
Arsenic	As	33	2	6	10	2	3	$[Ar]3d^{10}4s^24p^3$
Selenium	Se	34	2	6	10	2	4	$[Ar]3d^{10}4s^24p^4$
Bromine	Br	35	2	6	10	2	5	$[Ar]3d^{10}4s^24p^5$
Krypton	Kr	36	2	6	10	2	6	$[Ar]3d^{10}4s^24p^6$

*$[Ar] = 1s^22s^22p^63s^23p^6$

Elements of the Fifth Period

In the 18 elements of the fifth period, sublevels fill in a similar manner as in elements of the fourth period. However, they start at the 5s orbital instead of at the 4s orbital. Successive electrons are added first to the 5s orbital, then to the 4d orbitals, and finally to the 5p orbitals. This can be seen in **Figure 3.8.** There are occasional deviations from the predicted configurations here also. The deviations differ from those for fourth-period elements, but in each case the preferred configuration has the lowest possible energy.

FIGURE 3.8

ELECTRON CONFIGURATIONS OF ATOMS OF ELEMENTS IN THE FIFTH PERIOD

| Name | Symbol | Atomic number | Number of electrons in sublevels above 3d | | | | | Noble-gas notation |
			4s	4p	4d	5s	5p	
Rubidium	Rb	37	2	6		1		*[Kr]5s^1
Strontium	Sr	38	2	6		2		[Kr]5s^2
Yttrium	Y	39	2	6	1	2		[Kr]4$d^1$5s^2
Zirconium	Zr	40	2	6	2	2		[Kr]4$d^2$5s^2
Niobium	Nb	41	2	6	4	1		[Kr]4$d^4$5s^1
Molybdenum	Mo	42	2	6	5	1		[Kr]4$d^5$5s^1
Technetium	Tc	43	2	6	6	1		[Kr]4$d^6$5s^1
Ruthenium	Ru	44	2	6	7	1		[Kr]4$d^7$5s^1
Rhodium	Rh	45	2	6	8	1		[Kr]4$d^8$5s^1
Palladium	Pd	46	2	6	10			[Kr]4d^{10}
Silver	Ag	47	2	6	10	1		[Kr]4d^{10}5s^1
Cadmium	Cd	48	2	6	10	2		[Kr]4d^{10}5s^2
Indium	In	49	2	6	10	2	1	[Kr]4d^{10}5$s^2$5p^1
Tin	Sn	50	2	6	10	2	2	[Kr]4d^{10}5$s^2$5p^2
Antimony	Sb	51	2	6	10	2	3	[Kr]4d^{10}5$s^2$5p^3
Tellurium	Te	52	2	6	10	2	4	[Kr]4d^{10}5$s^2$5p^4
Iodine	I	53	2	6	10	2	5	[Kr]4d^{10}5$s^2$5p^5
Xenon	Xe	54	2	6	10	2	6	[Kr]4d^{10}5$s^2$5p^6

*[Kr] = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$

Electron Configurations

Sample Problem B

a. Write both the complete electron-configuration notation and the noble-gas notation for iron, Fe.

b. How many electron-containing orbitals are in an atom of iron? How many of these orbitals are completely filled? How many unpaired electrons are there in an atom of iron? In which sublevel are the unpaired electrons located?

● **SOLVE**

a. The complete electron-configuration notation of iron is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$. The periodic table inside the back cover of the text reveals that $1s^2 2s^2 2p^6 3s^2 3p^6$ is the electron configuration of the noble gas argon, Ar. Therefore, as shown in **Figure 3.7**, iron's noble-gas notation is $[Ar]3d^6 4s^2$.

b. An iron atom has 15 orbitals that contain electrons. They consist of one 1s orbital, one 2s orbital, three 2p orbitals, one 3s orbital, three 3p orbitals, five 3d orbitals, and one 4s orbital. Eleven of these orbitals are filled, and there are four unpaired electrons. They are located in the 3d sublevel. The notation $3d^6$ represents:

3d $\uparrow\downarrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$

Practice   Answers in Appendix E

1. a. Write both the complete electron-configuration notation and the noble-gas notation for iodine, I. How many inner-shell electrons does an iodine atom contain?

b. How many electron-containing orbitals are in an atom of iodine? How many of these orbitals are filled? How many unpaired electrons are there in an atom of iodine?

2. a. Write the noble-gas notation for tin, Sn. How many unpaired electrons are there in an atom of tin?

b. How many electron-containing d orbitals are there in an atom of tin? Name the element in the fourth period whose atoms have the same number of electrons in their highest energy levels that tin's atoms do.

3. a. Write the complete electron configuration for the element with atomic number 25. You may use the diagram shown in **Figure 3.2.**

b. Identify the element described in item 3a.

4. a. How many orbitals are completely filled in an atom of the element with atomic number 18? Write the complete electron configuration for this element.

b. Identify the element described in item 4a.

Elements of the Sixth Period

The sixth period consists of 32 elements. It is much longer than the periods that precede it in the periodic table. To build up electron configurations for elements of this period, electrons are added first to the 6s orbital in cesium, Cs, and barium, Ba. Then, in lanthanum, La, an electron is added to the 5d orbital.

With the next element, cerium, Ce, the $4f$ orbitals begin to fill, giving cerium atoms a configuration of $[Xe]4f^15d^16s^2$. In the next 13 elements, the $4f$ orbitals are filled. Next the $5d$ orbitals are filled, and the period is completed by filling the $6p$ orbitals. Because the $4f$ and the $5d$ orbitals are very close in energy, numerous deviations from the simple rules occur as these orbitals are filled. The electron configurations of the sixth-period elements can be found in the periodic table inside the back cover of the text. The seventh period is incomplete and consists largely of synthetic elements.

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Electron Configurations

Sample Problem C

a. Write both the complete electron-configuration notation and the noble-gas notation for a iridium atom.

b. Identify the elements in the second, third, and fourth periods that have the same number of highest-energy-level electrons as iridium.

SOLUTION

a. $1s^22s^2p^63s^2p^6d^{10}4s^2p^6d^{10}f^{14}5s^2p^6d^76s^2$; $[Xe]4f^{14}5d^76s^2$

b. Iridium has two electrons in its highest energy level (6). Elements with the same outermost configuration in their highest level are, in the second period, beryllium; in the third period, magnesium; in the fourth period, calcium, Ca; scandium, Sc; titanium, Ti; vanadium, Va; manganese, Mn; iron, Fe; cobalt, Co; nickel, Ni; zinc, Zn; germanium, Ge.

Practice   Answers in Appendix E

1. a. Write both the complete electron-configuration notation and the noble-gas notation for a barium atom.

   b. Identify the elements in the second, third, fourth, and fifth periods that have the same number of highest-energy-level electrons as barium.

2. a. Write the noble-gas notation for an antimony atom.

   b. Identify the elements in the sixth period that have an unpaired $5p$ electron.

✔ SECTION 3 FORMATIVE ASSESSMENT

▶ Reviewing Main Ideas

1. a. What is an atom's electron configuration?

   b. What three principles guide the electron configuration of an atom?

2. What three methods are used to represent the arrangement of electrons in atoms?

3. What is an octet of electrons? Which elements contain an octet of electrons?

4. Write the complete electron-configuration notation, the noble-gas notation, and the orbital notation for the following elements:

   a. carbon   **b.** neon   **c.** sulfur

5. Identify the elements having the following electron configurations:

   a. $1s^22s^22p^63s^23p^3$

   b. $[Ar]4s^1$

   c. contains just four electrons in its 3p orbitals

   d. contains one set of paired and three unpaired electrons in its fourth and outer main energy level

✔ Critical Thinking

6. RELATING IDEAS Write the electron configuration for the third-period elements Al, Si, P, S, and Cl. Is there a relationship between an element's group number and number of electrons in the outermost energy level?

Math Tutor

Weighted Averages and Atomic Mass

You have learned that the mass of a proton is about 1 u and that a neutron is only slightly heavier. Because atomic nuclei consist of whole numbers of protons and neutrons, you might expect that the atomic mass of an element would be very near a whole number. However, if you look at the periodic table, you will see that the atomic masses of many elements lie somewhere between whole numbers. In fact, the atomic masses listed on the table are *average* atomic masses. The atomic masses are averages because most elements occur in nature as a specific mixture of isotopes. For example, 75.76% of chlorine atoms have a mass of 34.969 u, and 24.24% have a mass of 36.966 u. If the isotopes were in a 1:1 ratio, you could simply add the masses of the two isotopes together and divide by 2. However, to account for the differing abundance of the isotopes, you must calculate a *weighted average*. For chlorine, the weighted average is 35.45 u. The following two examples demonstrate how weighted averages are calculated.

Sample Problem

A sample of naturally occurring silver consists of 51.839% Ag-107 (atomic mass 106.905 093 u) and 48.161% Ag-109 (atomic mass 108.904 756 u). What is the average atomic mass of silver?

To find average atomic mass, convert each percentage to a decimal equivalent and multiply by the atomic mass of the isotope.

$$0.518\ 39 \times 106.905\ 093\ u = 55.419\ u$$
$$\underline{0.481\ 61 \times 108.904\ 756\ u = 52.450\ u}$$
$$107.869\ u$$

Adding the masses contributed by each isotope gives an average atomic mass of 107.869 u. Note that this value for the average atomic mass of silver is very near the one given in the periodic table.

A sample of naturally occurring magnesium consists of 78.99% Mg-24 (atomic mass 23.985 042 u), 10.00% Mg-25 (atomic mass 24.985 837 u), and 11.01% Mg-26 (atomic mass 25.982 593 u). What is the average atomic mass of magnesium?

Again, convert each percentage to a decimal and multiply by the atomic mass of the isotope to get the mass contributed by each isotope.

$$0.7899 \times 23.985\ 042\ u = 18.95\ u$$
$$0.1000 \times 24.985\ 837\ u = \ \ 2.499\ u$$
$$\underline{0.1101 \times 25.982\ 593\ u = \ \ 2.861\ u}$$
$$24.31\ u$$

Adding the masses contributed by each isotope gives an average atomic mass of 24.31 u.

Practice

1. Rubidium occurs naturally as a mixture of two isotopes, 72.17% Rb-85 (atomic mass 84.911 792 u) and 27.83% Rb-87 (atomic mass 86.909 186 u). What is the average atomic mass of rubidium?

2. The element silicon occurs as a mixture of three isotopes: 92.22% Si-28, 4.69% Si-29, and 3.09% Si-30. The atomic masses of these three isotopes are as follows: Si-28 = 27.976 926 u, Si-29 = 28.976 495 u, and Si-30 = 29.973 770 u.

   Find the average atomic mass of silicon.

Summary

BIG IDEA An atom's nucleus is surrounded by electrons that have both a wave and a particle-like behavior.

SECTION 1 The Development of a New Atomic Model

KEY TERMS

- In the early twentieth century, light was determined to have a dual wave-particle nature.
- Quantum theory was developed to explain observations such as the photoelectric effect and the line-emission spectrum of hydrogen.
- Quantum theory states that electrons can exist only at specific atomic energy levels.
- When an electron moves from one main energy level to a main energy level of lower energy, a photon is emitted. The photon's energy equals the energy difference between the two levels.
- An electron in an atom can move from one main energy level to a higher main energy level only by absorbing an amount of energy exactly equal to the difference between the two levels.

electromagnetic radiation
electromagnetic spectrum
wavelength
frequency
photoelectric effect
quantum
photon
ground state
excited state
line-emission spectrum
continuous spectrum

SECTION 2 The Quantum Model of the Atom

KEY TERMS

- In the early twentieth century, electrons were determined to have a dual wave-particle nature.
- The Heisenberg uncertainty principle states that it is impossible to determine simultaneously the position and velocity of an electron or any other particle.
- Quantization of electron energies is a natural outcome of the Schrödinger wave equation, which describes the properties of an atom's electrons.
- An orbital, a three-dimensional region around the nucleus, shows the region in space where an electron is most likely to be found.
- The four quantum numbers that describe the properties of electrons in atomic orbitals are the principal quantum number, the angular momentum quantum number, the magnetic quantum number, and the spin quantum number.

Heisenberg uncertainty principle
quantum theory
orbital
quantum number
principal quantum number
angular momentum quantum number
magnetic quantum number
spin quantum number

SECTION 3 Electron Configurations

KEY TERMS

- The ground-state electron configuration of an atom can be written by using the Aufbau principle, Hund's rule, and the Pauli exclusion principle.
- Electron configurations can be depicted by using different types of notation. In this book, three types of notation are used: orbital notation, electron-configuration notation, and noble-gas notation.
- Electron configurations of some atoms, such as chromium, deviate from the predictions of the Aufbau principle, but the ground-state configuration that results is the configuration with the minimum possible energy.

electron configuration
Aufbau principle
Pauli exclusion principle
Hund's rule
noble gas
noble-gas configuration

CHAPTER 4 **Review**

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SECTION 1
The Development of a New Atomic Model

▶ **REVIEWING MAIN IDEAS**

1. **a.** List five examples of electromagnetic radiation.
   b. What is the speed of all forms of electromagnetic radiation in a vacuum?

2. Prepare a two-column table. List the properties of light that can best be explained by the wave theory in one column. List those best explained by the particle theory in the second column. You may want to consult a physics textbook for reference.

3. What are the frequency and wavelength ranges of visible light?

4. List the colors of light in the visible spectrum in order of increasing frequency.

5. In the early twentieth century, what two experiments involving light and matter could not be explained by the wave theory of light?

6. **a.** How are the wavelength and frequency of electromagnetic radiation related?
   b. How are the energy and frequency of electromagnetic radiation related?
   c. How are the energy and wavelength of electromagnetic radiation related?

7. Which theory of light—the wave or particle theory—best explains the following phenomena?
   a. the interference of light
   b. the photoelectric effect
   c. the emission of electromagnetic radiation by an excited atom

8. Distinguish between the ground state and an excited state of an atom.

9. According to Bohr's model, how is hydrogen's emission spectrum produced?

PRACTICE PROBLEMS

10. Determine the frequency of light whose wavelength is 4.257×10^{-7} cm.

11. Determine the energy in joules of a photon whose frequency is 3.55×10^{17} Hz.

12. Using the two equations $E = h\nu$ and $c = \lambda\nu$, derive an equation expressing E in terms of h, c, and λ.

13. How long would it take a radio wave whose frequency is 7.25×10^5 Hz to travel from Mars to Earth if the distance between the two planets is approximately 8.00×10^7 km?

14. Cobalt-60 is an artificial radioisotope that is produced in a nuclear reactor and is used as a gamma-ray source in the treatment of certain types of cancer. If the wavelength of the gamma radiation from a cobalt-60 source is 1.00×10^{-3} nm, calculate the energy of a photon of this radiation.

SECTION 2
The Quantum Model of the Atom

▶ **REVIEWING MAIN IDEAS**

15. Describe two major shortcomings of Bohr's model of the atom.

16. **a.** What is the principal quantum number?
    b. How is it symbolized?
    c. What are shells?
    d. How does n relate to the number of electrons allowed per main energy level?

17. **a.** What information is given by the angular momentum quantum number?
    b. What are sublevels, or subshells?

18. For each of the following values of n, indicate the numbers and types of sublevels possible for that main energy level. (Hint: See **Figure 2.6.**)
    a. $n = 1$
    b. $n = 2$
    c. $n = 3$
    d. $n = 4$
    e. $n = 7$ (number only)

19. **a.** What information is given by the magnetic quantum number?
    b. How many orbital orientations are possible in each of the s, p, d, and f sublevels?
    c. Explain and illustrate the notation for distinguishing between the different p orbitals in a sublevel.

20. a. What is the relationship between n and the total number of orbitals in a main energy level?
b. How many total orbitals are contained in the third main energy level? in the fifth?

21. a. What information is given by the spin quantum number?
b. What are the possible values for this quantum number?

22. How many electrons could be contained in the following main energy levels with n equal to the number provided?
a. 1
b. 3
c. 4
d. 6
e. 7

PRACTICE PROBLEMS

23. Sketch the shape of an s orbital and a p orbital.

24. How does a $2s$ orbital differ from a $1s$ orbital?

25. How do a $2p_x$ and a $2p_y$ orbital differ?

Electron Configurations

▶ **REVIEWING MAIN IDEAS**

26. a. In your own words, state the Aufbau principle.
b. Explain the meaning of this principle in terms of an atom with many electrons.

27. a. In your own words, state Hund's rule.
b. What is the basis for this rule?

28. a. In your own words, state the Pauli exclusion principle.
b. What is the significance of the spin quantum number?

29. a. What is meant by the highest occupied energy level in an atom?
b. What are inner-shell electrons?

30. Determine the highest occupied energy level in the following elements:
a. He
b. Be
c. Al
d. Ca
e. Sn

31. Write the orbital notation for the following elements. (Hint: See Sample Problem A.)
a. P
b. B
c. Na
d. O

32. Write the electron-configuration notation for the element whose atoms contain the following number of electrons:
a. 3
b. 6
c. 8
d. 13

33. Given that the electron configuration for oxygen is $1s^2 2s^2 2p^4$, answer the following questions:
a. How many electrons are in each oxygen atom?
b. What is the atomic number of this element?
c. Write the orbital notation for oxygen's electron configuration.
d. How many unpaired electrons does oxygen have?
e. What is the highest occupied energy level?
f. How many inner-shell electrons does the atom contain?
g. In which orbital(s) are these inner-shell electrons located?

34. a. What are the noble gases?
b. What is a noble-gas configuration?
c. How does noble-gas notation simplify writing an atom's electron configuration?

35. Write the noble-gas notation for the electron configuration of each of the elements below. (Hint: See Sample Problem B.)
a. Cl
b. Ca
c. Se

36. a. What information is given by the noble-gas notation $[Ne]3s^2$?
b. What element does this represent?

37. Write both the complete electron-configuration notation and the noble-gas notation for each of the elements below. (Hint: See Sample Problem C.)
a. Na
b. Sr
c. P

38. Identify each of the following atoms on the basis of its electron configuration:
  a. $1s^22s^22p^1$
  b. $1s^22s^22p^5$
  c. $[Ne]3s^2$
  d. $[Ne]3s^23p^2$
  e. $[Ne]3s^23p^5$
  f. $[Ar]4s^1$
  g. $[Ar]3d^64s^2$

PRACTICE PROBLEMS

39. List the order in which orbitals generally fill, from the $1s$ to the $7p$ orbital.

40. Write the noble-gas notation for the electron configurations of each of the following elements:
  a. As   **e.** Sn
  b. Pb   **f.** Xe
  c. Lr   **g.** La
  d. Hg

41. How do the electron configurations of chromium and copper contradict the Aufbau principle?

Mixed Review

▶ REVIEWING MAIN IDEAS

42. a. Which has a longer wavelength: green light or yellow light?
  b. Which has a higher frequency: an x-ray or a microwave?
  c. Which travels at a greater speed: ultraviolet light or infrared light?

43. Write both the complete electron-configuration and noble-gas notation for each of the following:
  a. Ar   **b.** Br   **c.** Al

44. Given the speed of light as 3.00×10^8 m/s, calculate the wavelength of the electromagnetic radiation whose frequency is 7.500×10^{12} Hz.

45. a. What is the electromagnetic spectrum?
  b. What units can be used to express wavelength?
  c. What unit is used to express frequencies of electromagnetic waves?

46. Given that the electron configuration for phosphorus is $1s^22s^22p^63s^23p^3$, answer the following questions:
  a. How many electrons are in each atom?
  b. What is the atomic number of this element?
  c. Write the orbital notation for this element.
  d. How many unpaired electrons does an atom of phosphorus have?
  e. What is its highest occupied energy level?
  f. How many inner-shell electrons does the atom contain?
  g. In which orbital(s) are these inner-shell electrons located?

47. What is the frequency of a radio wave whose energy is 1.55×10^{-24} J per photon?

48. Write the noble-gas notation for the electron configurations of each of the following elements:
  a. Hf   **d.** At
  b. Sc   **e.** Ac
  c. Fe   **f.** Zn

49. Describe the major similarities and differences between Schrödinger's model of the atom and the model proposed by Bohr.

50. When sodium is heated, a yellow spectral line whose energy is 3.37×10^{-19} J per photon is produced.
  a. What is the frequency of this light?
  b. What is the wavelength of this light?

51. a. What is an orbital?
  b. Describe an orbital in terms of an electron cloud.

CRITICAL THINKING

52. Inferring Relationships In the emission spectrum of hydrogen shown in **Figure 1.5**, each colored line is produced by the emission of photons with specific energies. Substances also produce absorption spectra when electromagnetic radiation passes through them. Certain wavelengths are absorbed. Using the diagram below, predict what the wavelengths of the absorption lines will be when white light (all of the colors of the visible spectrum) is passed through hydrogen gas.

300 nm   700 nm
Hydrogen absorption spectrum
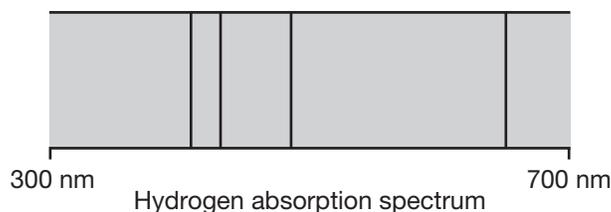
53. Applying Models In discussions of the photoelectric effect, the minimum energy needed to remove an electron from the metal is called the *threshold energy* and is a characteristic of the metal. For example, chromium, Cr, will emit electrons when the wavelength of the radiation is 284 nm or less. Calculate the threshold energy for chromium. (Hint: You will need to use the two equations that describe the relationships between wavelength, frequency, speed of light, and Planck's constant.)

54. Analyzing Information Four electrons in an atom have the four sets of quantum numbers given below. Which electrons are in the same orbital? Explain your answer.
a. 1, 0, 0, $-1/2$
b. 1, 0, 0, $+1/2$
c. 2, 1, 1, $+1/2$
d. 2, 1, 0, $+1/2$

55. Relating Ideas Which of the sets of quantum numbers below are possible? Which are impossible? Explain your choices.
a. 2, 2, 1, $+1/2$
b. 2, 0, 0, $-1/2$
c. 2, 0, 1, $-1/2$

USING THE HANDBOOK

56. Sections 1 and 2 of the *Elements Handbook* (Appendix A) contain information on an analytical test and a technological application for Group 1 and 2 elements. The test and application are based on the emission of light from atoms. Review these sections to answer the following:
a. What analytical technique utilizes the emission of light from excited atoms?
b. What elements in Groups 1 and 2 can be identified by this technique?
c. What types of compounds are used to provide color in fireworks?
d. What wavelengths within the visible spectrum would most likely contain emission lines for barium?

RESEARCH AND WRITING

57. Neon signs do not always contain neon gas. The various colored lights produced by the signs are due to the emission of a variety of low-pressure gases in different tubes. Research other kinds of gases used in neon signs, and list the colors that they emit.

58. Prepare a report about the photoelectric effect, and cite some of its practical uses. Explain the basic operation of each device or technique mentioned.

ALTERNATIVE ASSESSMENT

59. Performance A spectroscope is a device used to produce and analyze spectra. Construct a simple spectroscope, and determine the absorption spectra of several elemental gases. (Your teacher will provide you with the gas discharge tubes containing samples of different gases.)

Standards-Based Assessment

Record your answers on a separate piece of paper.

MULTIPLE CHOICE

1 For which of the following discoveries is Niels Bohr given credit?

  A discovering the electron
  B linking spectral lines with electron energy levels
  C explaining the details of bioluminescence
  D explaining the chemical behavior of atoms

Light	wavelength	frequency
Red	6.5×10^{-7} m	?
Blue	?	6.60×10^{14} Hz

speed of light	3×10^8 m/s
Planck's constant	6.26×10^{-34} J·s

For questions 2–4, refer to the tables above.

2 Using the table above, calculate the frequency of red light.

  A 4.6×10^{14} Hz
  B 3.0×10^8 Hz
  C 1.3×10^6 Hz
  D 9.8×10^4 Hz

3 Using the table above, calculate the wavelength of blue light.

  A 3.9×10^{-3} m
  B 2.2×10^{-7} m
  C 3.0×10^{-8} m
  D 4.5×10^{-7} m

4 How much energy in joules is found in a photon of yellow light with wavelength 5.9×10^{-7} m?

  A 1.3×10^{14} J
  B 3.9×10^{-19} J
  C 3.4×10^{-19} J
  D 3.4×10^{14} J

5 Which element has the noble gas notation [Kr]$5s^24d^2$?

  A Se
  B Zr
  C Sr
  D Mo

6 When a calcium salt is heated in a flame, a photon of light with an energy of 3.2×10^{-19} J is emitted. On the basis of this fact and the table below, what color would be expected for the calcium flame?

Frequency, s^{-1}	7.1×10^{14}	6.4×10^{14}	5.7×10^{14}
Wavelength, nm	422	469	526
Color	violet	blue	green
Frequency, s^{-1}	5.2×10^{14}	4.8×10^{14}	4.3×10^{14}
Wavelength, nm	577	625	698
Color	yellow	orange	red

  A yellow
  B yellow-orange
  C orange
  D red

7 Which of the following correctly expresses the electron configuration for bromine, Br?

  A $1s^22s^22p^63s^23p^63d^{10}4s^24p^5$
  B $1s^22s^22p^63s^23p^63d^84s^24p^63d^84s^24p^5$
  C $1s^22p^63s^23p^63d^{10}4s^24p^63d^{10}4s^24p^5$
  D $1s^22s^22p^63s^23p^63d^{10}4s^24p^63d^{10}4s^24p^6$

GRIDDED RESPONSE

8 A sample of a naturally occurring element contains the following proportions:

Natural Abundance	Atomic Mass
19.9%	10.0129369
80.1%	11.0093054

What is the average atomic mass for this element? Round to the nearest thousandth.

Test Tip
If time permits, take short mental breaks during the test to improve your concentration.