5

Electrons in Atoms

INSIDE:

- **5.1** Revising the Atomic Model
- **5.2** Electron Arrangement in Atoms
- **5.3** Atomic Emission Spectra and the Quantum Mechanical Model

PearsonChem.com

CONCEPTS IN ACTION

KINETIC ART

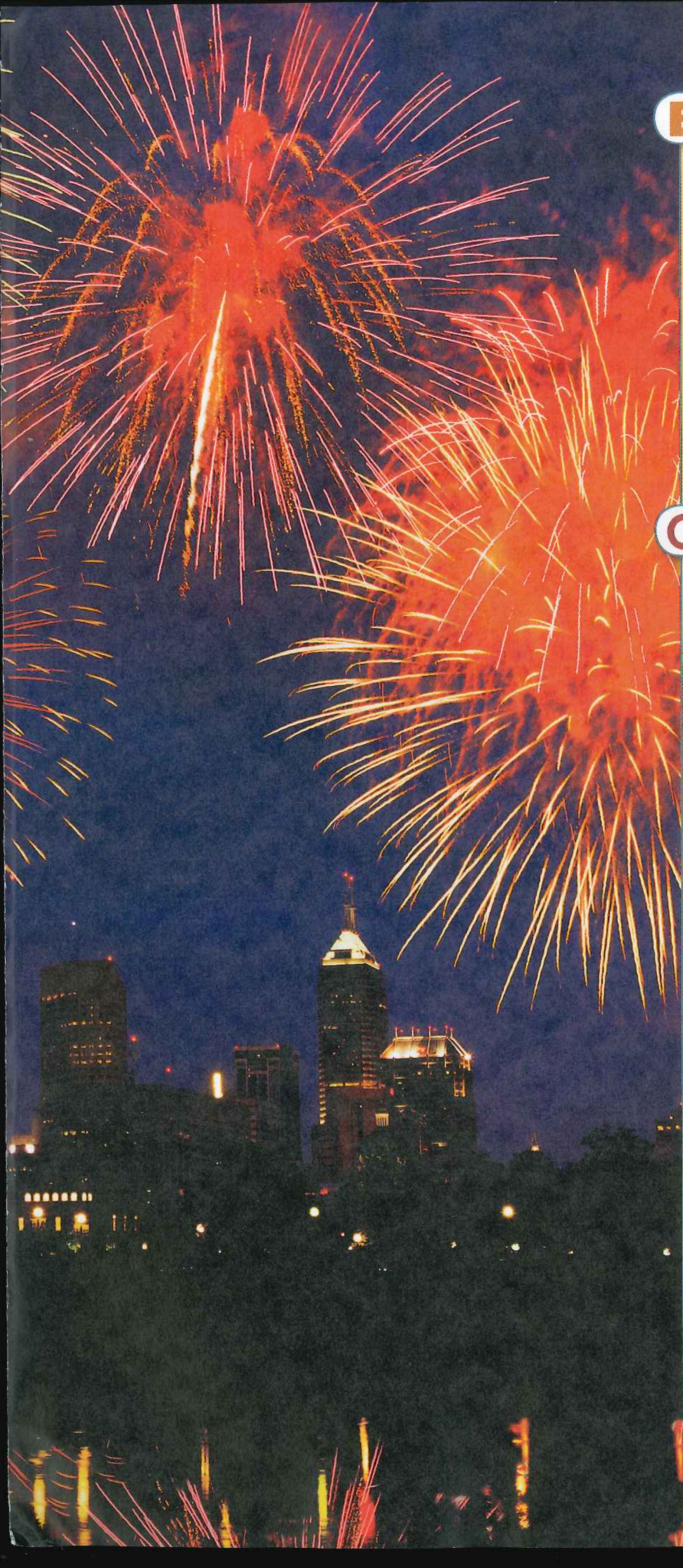
VIRTUAL LAB

MATH TUTOR

ONLINE PROBLEMS

CHEM TUTOR

The brilliant colors of fireworks are produced by using compounds containing different elements. In this chapter, you will learn how elements can emit light of different colors.



BIGIDEA

ELECTRONS AND THE STRUCTURE OF ATOMS

Essential Questions:

1. *How does the quantum mechanical model describe the arrangement of electrons in atoms?*
2. *What happens when electrons in atoms absorb or release energy?*

CHEMYSTERY

Now You See It… Now You Don't

Young Liam loves to go outside at night and gaze at the stars. For his birthday, his parents gave him glow-in-the-dark star stickers so that he could look at the stars from the comfort of his bed.

That night, Liam looked at the constellations he had created on his ceiling while he drifted off to sleep. A few hours later, Liam woke up very disturbed. The stars were no longer glowing even though it was still dark in his room. He turned on his bedroom light and ran down the hall to wake his parents. However, when Liam brought his parents back to his room and turned off the light, the stars were glowing again. Why did the stars cease to glow and then light up again later?

▶ Connect to the **BIGIDEA** As you read about electrons in atoms, think about how glow-in-the-dark stickers might work.

NATIONAL SCIENCE EDUCATION STANDARDS

A-1, A-2, B-1, B-6, E-2, G-1, G-2, G-3

5.1 Revising the Atomic Model

CHEMISTRY & YOU

Q: *Why do scientists use mathematical models to describe the position of electrons in atoms?* Wind tunnels and models are often used to simulate the forces from the moving air on a design. Shown here is a life-sized model of a speed skier. It is a physical model. However, not all models are physical. In fact, the current model of the atom is a mathematical model.

Key Questions

What did Bohr propose in his model of the atom?

What does the quantum mechanical model determine about the electrons in an atom?

How do sublevels of principal energy levels differ?

Vocabulary

- energy level
- quantum
- quantum mechanical model
- atomic orbital

Energy Levels in Atoms

What did Bohr propose in his model of the atom?

Thus far, the atomic model presented in this textbook has considered atoms as consisting of protons and neutrons making up a nucleus surrounded by electrons. After discovering the atomic nucleus, Rutherford used existing ideas about the atom and proposed an atomic model in which the electrons move around the nucleus like the planets move around the sun.

Limitations of Rutherford's Atomic Model Rutherford's atomic model explained only a few simple properties of atoms. It could not explain the chemical properties of elements. For example, Rutherford's model could not explain why metals or compounds of metals give off characteristic colors when heated in a flame. It also could not explain why an object such as the iron scroll shown in Figure 5.1 first glows dull red, then yellow, and then white when heated to higher and higher temperatures. Explaining what leads to the chemical properties of elements required a model that better described the behavior of electrons in atoms.

Figure 5.1 Glowing Metal
Rutherford's model failed to explain why objects change color when heated. As the temperature of this iron scroll is increased, it first appears black, then red, then yellow, and then white. The observed behavior could be explained only if the atoms in the iron gave off light in specific amounts of energy. A better atomic model was needed to explain this observation.

The Bohr Model In 1913, Niels Bohr (1885–1962), a young Danish physicist and a student of Rutherford, developed a new atomic model. He changed Rutherford's model to incorporate newer discoveries about how the energy of an atom changes when the atom absorbs or emits light. He considered the simplest atom, hydrogen, which has one electron. **Bohr proposed that an electron is found only in specific circular paths, or orbits, around the nucleus.**

Each possible electron orbit in Bohr's model has a fixed energy. The fixed energies an electron can have are called **energy levels.** The fixed energy levels of electrons are somewhat like the rungs of the ladder in Figure 5.2a. The lowest rung of the ladder corresponds to the lowest energy level. A person can climb up or down the ladder by stepping from rung to rung. Similarly, an electron can move from one energy level to another. A person on the ladder cannot stand between the rungs. Similarly, the electrons in an atom cannot exist between energy levels. To move from one rung to another, a person climbing the ladder must move just the right distance. To move from one energy level to another, an electron must gain or lose just the right amount of energy. A **quantum** of energy is the amount of energy required to move an electron from one energy level to another energy level. The energy of an electron is therefore said to be quantized.

READING SUPPORT

Build Vocabulary: *Latin Word Origins* *Quantum* comes from the Latin word *quantus*, meaning "how much." ***What other commonly used English word comes from this root?***

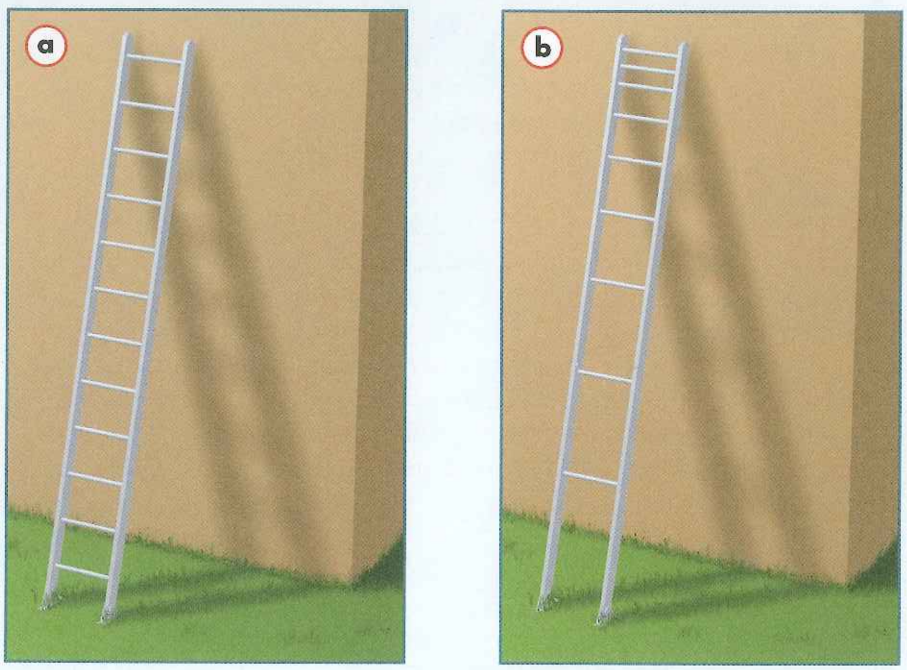


Figure 5.2 Energy Levels
The rungs of a ladder are somewhat like the energy levels in Bohr's model of the atom. **a.** In an ordinary ladder, the rungs are equally spaced. **b.** The energy levels in atoms are unequally spaced, like the rungs in this unusual ladder. The higher energy levels are closer together.
Compare ***For the ladder in b, compare the amount of energy it would take to move from the first rung to the second rung with the amount of energy it would take to move from the second rung to the third rung.***

The amount of energy an electron gains or loses in an atom is not always the same. Like the rungs of the strange ladder in Figure 5.2b, the energy levels in an atom are not equally spaced. The higher energy levels are closer together. It takes less energy to climb from one rung to another near the top of the ladder in Figure 5.2b, where the rungs are closer. Similarly, the higher the energy level occupied by an electron, the less energy it takes the electron to move from that energy level to the next higher energy level.

The Bohr model provided results in agreement with experiments using the hydrogen atom. However, the Bohr model failed to explain the energies absorbed and emitted by atoms with more than one electron.

The Quantum Mechanical Model

What does the quantum mechanical model determine about the electrons in an atom?

The Rutherford model and the Bohr model of the atom described the path of a moving electron as you would describe the path of a large moving object. Later theoretical calculations and experimental results were inconsistent with describing electron motion this way. In 1926, the Austrian physicist Erwin Schrödinger (1887–1961) used these calculations and results to devise and solve a mathematical equation describing the behavior of the electron in a hydrogen atom. The modern description of the electrons in atoms, the **quantum mechanical model,** came from the mathematical solutions to the Schrödinger equation.

Like the Bohr model, the quantum mechanical model of the atom restricts the energy of electrons to certain values. Unlike the Bohr model, however, the quantum mechanical model does not specify an exact path the electron takes around the nucleus. **The quantum mechanical model determines the allowed energies an electron can have and how likely it is to find the electron in various locations around the nucleus of an atom.**

Probability describes how likely it is to find an electron in a particular location around the nucleus of an atom. If you placed three red marbles and one green marble into a box and then picked a marble without looking, the probability of picking the green marble would be one in four, or 25 percent. This percentage means that if you put the four marbles in a box and picked one, and repeated this many times, you would pick a green marble in 25 percent of your tries.

The quantum mechanical model description of how electrons move around the nucleus is similar to a description of how the blades of a windmill rotate. The windmill blades in Figure 5.3a have some probability of being anywhere in the blurry region they produce in the picture, but you cannot predict their exact locations at any instant. In the quantum mechanical model of the atom, the probability of finding an electron within a certain volume of space surrounding the nucleus can be represented as a fuzzy cloudlike region, as shown in Figure 5.3b. The cloud is more dense where the probability of finding the electron is high and is less dense where the probability of finding the electron is low. There is no boundary to the cloud because there is a slight chance of finding the electron at a considerable distance from the nucleus. Therefore, attempts to show probabilities as a fuzzy cloud are usually limited to the volume in which the electron is found 90 percent of the time. To visualize an electron probability cloud, imagine that you could mold a sack around the cloud so that the electron was inside the sack 90 percent of the time. The shape of the sack would then give you a picture of the shape of the cloud.

Figure 5.3 Electron Cloud
The electron cloud of an atom can be compared to a photograph of spinning windmill blades. **a.** The windmill blades are somewhere in the blurry region they produce in this picture, but the picture does not tell you their exact positions at any instant. **b.** Similarly, the electron cloud of an atom represents the locations where an electron is likely to be found, but it is not possible to know where an electron is in the cloud at any instant.

Atomic Orbitals

How do sublevels of principal energy levels differ?

Solutions to the Schrödinger equation give the energies, or energy levels, an electron can have. For each energy level, the Schrödinger equation also leads to a mathematical expression, called an **atomic orbital,** describing the probability of finding an electron at various locations around the nucleus. An atomic orbital is represented pictorially as a region of space in which there is a high probability of finding an electron.

The energy levels of electrons in the quantum mechanical model are labeled by principal quantum numbers (n). These numbers are assigned the values $n = 1, 2, 3, 4$, and so forth. For each principal energy level greater than 1, there are several orbitals with different shapes and at different energy levels. These energy levels within a principal energy level constitute energy sublevels. **Each energy sublevel corresponds to one or more orbitals of different shapes. The orbitals describe where an electron is likely to be found.**

Different atomic orbitals are denoted by letters. As shown in Figure 5.4a, *s* orbitals are spherical, and *p* orbitals are dumbbell-shaped. The probability of finding an electron at a given distance from the nucleus in an *s* orbital does not depend on direction because of its spherical shape. The three kinds of *p* orbitals have different orientations in space. Figure 5.4b shows the shapes of *d* orbitals. Four of the five kinds of *d* orbitals have cloverleaf shapes. The shapes of *f* orbitals are more complicated than the shapes of *d* orbitals.

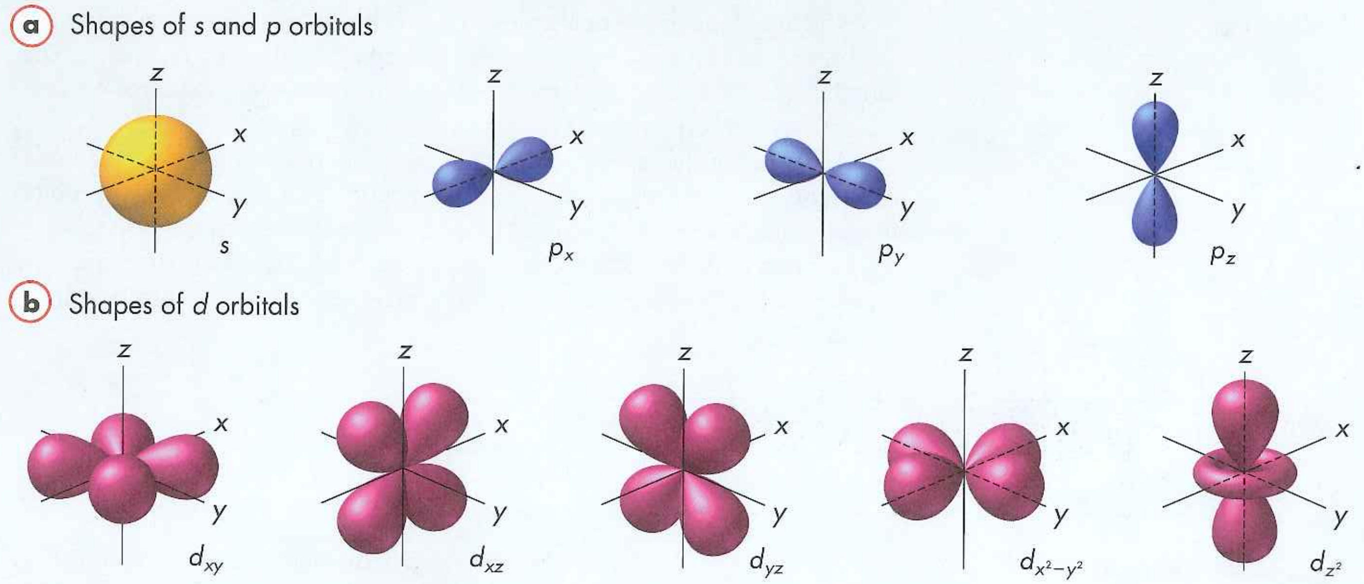


Figure 5.4 Atomic Orbitals
Solutions to the Schrödinger equation give rise to atomic orbitals. **a.** For a given principal energy level greater than 1, there is one *s* orbital and three *p* orbitals. **b.** Four of the five *d* orbitals have the same shape but different spatial orientations.
Interpret Diagrams *How are the orientations of the d_{xy} and $d_{x^2-y^2}$ orbitals similar? How are they different?*

See atomic orbitals animated online. KINETIC ART

CONCEPTS IN ACTION Learn more about probability and atomic structure online.

Table 5.1

Summary of Principal Energy Levels and Sublevels

Principal energy level	Number of sublevels	Type of sublevel	Maximum number of electrons
$n = 1$	1	1*s* (1 orbital)	2
$n = 2$	2	2*s* (1 orbital), 2*p* (3 orbitals)	8
$n = 3$	3	3*s* (1 orbital), 3*p* (3 orbitals), 3*d* (5 orbitals)	18
$n = 4$	4	4*s* (1 orbital), 4*p* (3 orbitals), 4*d* (5 orbitals), 4*f* (7 orbitals)	32

CHEMISTRY & YOU

Q: *Previous models of the atom were physical models based on the motion of large objects. Why do scientists no longer use physical models to describe the motion of electrons?*

As shown in Table 5.1, the numbers and types of atomic orbitals depend on the principal energy level. The lowest principal energy level ($n = 1$) has only one sublevel, called 1*s*. The second principal energy level ($n = 2$) has two sublevels, 2*s* and 2*p*. The 2*p* sublevel is of higher energy than the 2*s* sublevel and consists of three *p* orbitals of equal energy. Thus the second principal energy level has four orbitals (one 2*s* and three 2*p* orbitals).

The third principal energy level ($n = 3$) has three sublevels. These are called 3*s*, 3*p*, and 3*d*. The 3*d* sublevel consists of five *d* orbitals of equal energy. Thus the third principal energy level has nine orbitals (one 3*s*, three 3*p*, and five 3*d* orbitals).

The fourth principal energy level ($n = 4$) has four sublevels, called 4*s*, 4*p*, 4*d*, and 4*f*. The 4*f* sublevel consists of seven *f* orbitals of equal energy. The fourth principal energy level, then, has sixteen orbitals (one 4*s*, three 4*p*, five 4*d*, and seven 4*f* orbitals).

As shown in the table, the principal quantum number always equals the number of sublevels within that principal energy level. The number of orbitals in a principal energy level is equal to n^2. As you will learn in the next lesson, a maximum of two electrons can occupy an orbital. Therefore, the maximum number of electrons that can occupy a principal energy level is given by the formula $2n^2$.

ONLINE PROBLEMS 5.1 LessonCheck

1. **Review** What was the basic proposal in the Bohr model of the atom?
2. **Describe** What does the quantum mechanical model determine about electrons in atoms?
3. **Review** How do two sublevels of the same principal energy level differ from each other?
4. **Describe** How can electrons in an atom move from one energy level to another?
5. **Explain** The energies of electrons are said to be quantized. Explain what this means.
6. **Apply Concepts** How many orbitals are in the following sublevels?
   - **a.** 3*p* sublevel
   - **b.** 2*s* sublevel
   - **c.** 4*p* sublevel
   - **d.** 3*d* sublevel
   - **e.** 4*f* sublevel

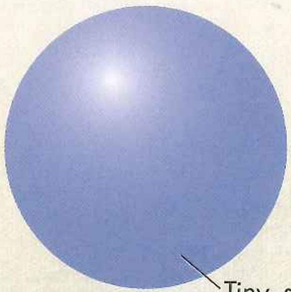
BIGIDEA

ELECTRONS AND THE STRUCTURE OF ATOMS

7. How do the Bohr model and the quantum mechanical model differ in the way they describe the arrangement of electrons in atoms?

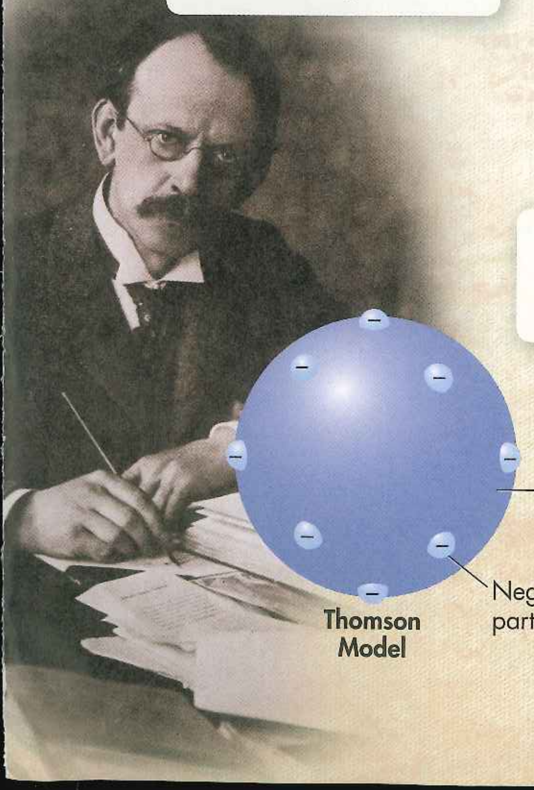
Development of Atomic Models

The atomic model has changed as scientists learned more about the atom's structure through experiments and calculations.



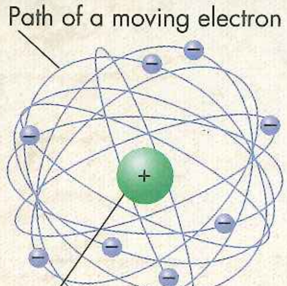
1803 John Dalton pictures atoms as tiny, indestructible particles.

1897 J.J. Thomson discovers the electron. He pictures electrons embedded in a sphere of positive electrical charge.

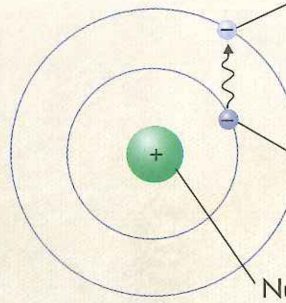


1904 Hantaro Nagaoka suggests that an atom has a central nucleus. Electrons move in orbits like the rings around Saturn.

1911 Ernest Rutherford finds that an atom has a small, dense, positively charged nucleus.

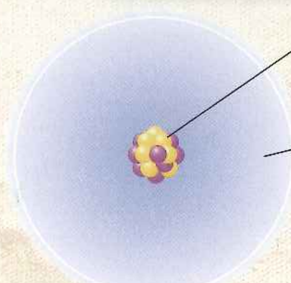


1913 In Niels Bohr's model, the electron moves in a circular orbit at fixed distances from the nucleus.



1923 Louis de Broglie proposes that moving particles like electrons have some properties of waves.

1926 Erwin Schrödinger develops mathematical equations to describe the motion of electrons in atoms, which leads to the quantum mechanical model.



1932 James Chadwick confirms the existence of neutrons, which have no charge. Atomic nuclei contain neutrons and positively charged protons.

1800 1805 | 1895 1900 1905 1910 1915 1920 1925 1930 1935

Take It Further

1. Summarize List a major contribution of each of these scientists to the understanding of the atom: Dalton, Thomson, Rutherford, Bohr, and Schrödinger.

2. Describe Have you ever needed to identify something you could not see? Explain.

5.2 Electron Arrangement in Atoms

CHEMISTRY & YOU

Q: *What makes the electron configuration of an atom stable?* Unstable arrangements, such as the yoga position shown here, tend to become more stable by losing energy. If the yogi were to fall, she would have less energy, but her position would be more stable. Energy and stability play an important role in determining how electrons are configured in an atom.

Key Question

What are the three rules for writing the electron configurations of elements?

Vocabulary

- electron configuration
- aufbau principle
- Pauli exclusion principle
- spin
- Hund's rule

Electron Configurations

What are the three rules for writing the electron configurations of elements?

In an atom, electrons and the nucleus interact to make the most stable arrangement possible. The ways in which electrons are arranged in various orbitals around the nuclei of atoms are called **electron configurations.** **Three rules—the aufbau principle, the Pauli exclusion principle, and Hund's rule—tell you how to find the electron configurations of atoms.** The three rules are as follows.

Aufbau Principle According to the **aufbau principle,** electrons occupy the orbitals of lowest energy first. The orbitals for any sublevel of a principal energy level are always of equal energy. Within a principal energy level, the *s* sublevel is always the lowest-energy sublevel. However, the range of energy levels within a principal energy level can overlap the energy levels of another principal level. Look at the aufbau diagram in Figure 5.5. Each box represents an atomic orbital. Notice that the filling of atomic orbitals does not follow a simple pattern beyond the second energy level. For example, the 4*s* orbital is lower in energy than a 3*d* orbital.

Pauli Exclusion Principle According to the **Pauli exclusion principle,** an atomic orbital may describe at most two electrons. For example, either one or two electrons can occupy an *s* orbital or a *p* orbital. To occupy the same orbital, two electrons must have opposite spins; that is, the electron spins must be paired. **Spin** is a quantum mechanical property of electrons and may be thought of as clockwise or counterclockwise. A vertical arrow indicates an electron and its direction of spin ($\uparrow$ or $\downarrow$). An orbital containing paired electrons is written as $\boxed{\uparrow\downarrow}$.

Hund's Rule According to **Hund's rule,** electrons occupy orbitals of the same energy in a way that makes the number of electrons with the same spin direction as large as possible. For example, three electrons would occupy three orbitals of equal energy as follows: $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$. Electrons then occupy each orbital so that their spins are paired with the first electron in the orbital.

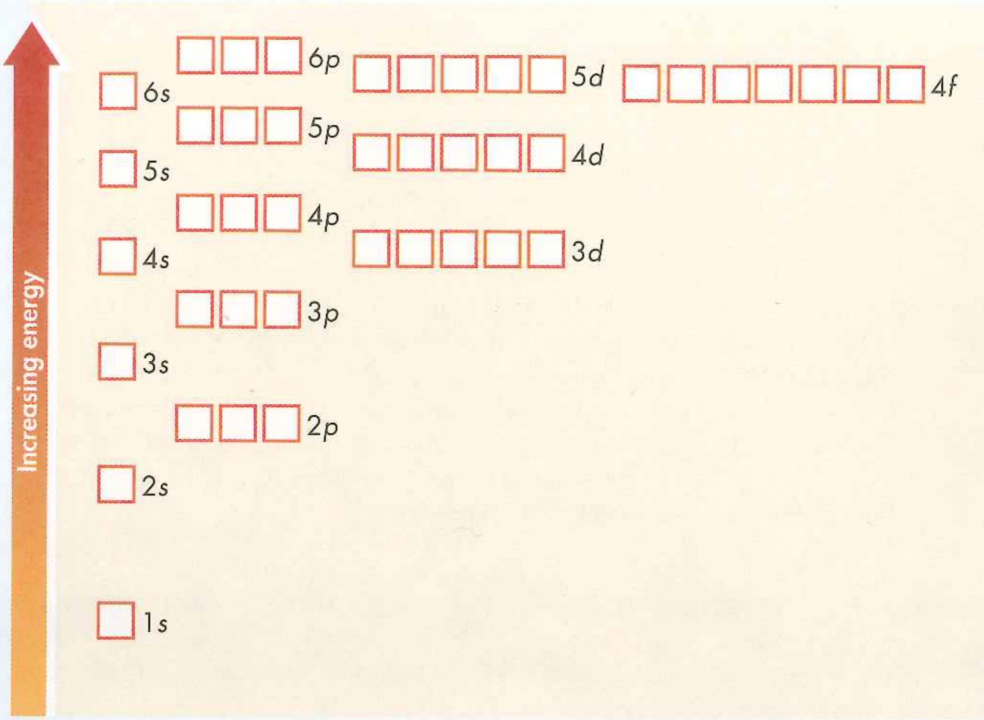


Figure 5.5 Aufbau Diagram
This aufbau diagram shows the relative energy levels of the various atomic orbitals. Orbitals of greater energy are higher on the diagram.
Interpret Tables ***Which is of higher energy, a 4d orbital or a 5s orbital?***

Look at the orbital filling diagrams of the atoms listed in Table 5.2. An oxygen atom contains eight electrons. The orbital of lowest energy, 1*s*, has one electron, then a second electron of opposite spin. The next orbital to fill is 2*s*. It also has one electron, then a second electron of opposite spin. One electron then occupies each of the three 2*p* orbitals of equal energy. The remaining electron now pairs with an electron occupying one of the 2*p* orbitals. The other two 2*p* orbitals remain only half filled, with one electron each.

Table 5.2

Electron Configurations of Selected Elements

Element	1s	2s	$2p_x$	$2p_y$	$2p_z$	3s	Electron configuration
H	↑						$1s^1$
He	↑↓						$1s^2$
Li	↑↓	↑					$1s^2 2s^1$
C	↑↓	↑↓	↑	↑			$1s^2 2s^2 2p^2$
N	↑↓	↑↓	↑	↑	↑		$1s^2 2s^2 2p^3$
O	↑↓	↑↓	↑↓	↑	↑		$1s^2 2s^2 2p^4$
F	↑↓	↑↓	↑↓	↑↓	↑		$1s^2 2s^2 2p^5$
Ne	↑↓	↑↓	↑↓	↑↓	↑↓		$1s^2 2s^2 2p^6$
Na	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$1s^2 2s^2 2p^6 3s^1$

CHEMISTRY & YOU

Q: *Explain why the correct electron configuration of oxygen is $1s^2 2s^2 2p^4$ and not $1s^2 2s^2 2p^3 3s^1$.*

A convenient shorthand method for showing the electron configuration of an atom involves writing the energy level and the symbol for every sublevel occupied by an electron. You indicate the number of electrons occupying each sublevel with a superscript. For hydrogen, with one electron in a 1*s* orbital, the electron configuration is written $1s^1$. For helium, with two electrons in a 1*s* orbital, the configuration is $1s^2$. For oxygen, with two electrons in a 1*s* orbital, two electrons in a 2*s* orbital, and four electrons in 2*p* orbitals, the electron configuration is $1s^2 2s^2 2p^4$. Note that the sum of the superscripts equals the number of electrons in the atom.

In this book, when electron configurations are written, the sublevels within the same principal energy level are generally written together. These configurations are not always in the same order as shown on the aufbau diagram. For example, the electron configuration of bromine is written as $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$. The 3*d* sublevel is written before the 4*s* sublevel, even though the 4*s* sublevel has lower energy.

CHEM TUTOR

Sample Problem 5.1

Writing Electron Configurations

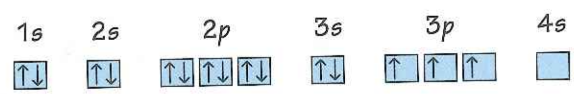
The atomic number of phosphorus is 15. Write the electron configuration of a phosphorus atom.

When writing electron configurations, the sublevels within the same principal energy level are written together.

1 Analyze **Identify the relevant concepts.** Phosphorus has 15 electrons. There is a maximum of two electrons per orbital. Electrons do not pair up within an energy sublevel (orbitals of equal energy) until each orbital already has one electron.

2 Solve **Apply the concepts to this problem.**

Use the aufbau diagram in Figure 5.5 to place electrons in the orbital with the lowest energy (1*s*) first. Continue placing electrons in each orbital with the next higher energy level.



Write the electron configuration.

The electron configuration of phosphorus is $1s^2 2s^2 2p^6 3s^2 3p^3$. The superscripts add up to the number of electrons.

8. Write the electron configuration for each atom.

a. carbon

b. argon

c. nickel

9. Write the electron configuration for each atom. How many unpaired electrons does each atom have?

a. boron

b. silicon

c. sulfur

Exceptional Electron Configurations Copper, which is shown in Figure 5.6, has an electron configuration that is an exception to the aufbau principle. You can obtain correct electron configurations for the elements up to vanadium (atomic number 23) by following the aufbau diagram for orbital filling. If you were to continue in that fashion, however, you would assign chromium and copper the following incorrect configurations.

$$\text{Cr } 1s^2 2s^2 2p^6 3s^2 3p^6 3d^4 4s^2$$
$$\text{Cu } 1s^2 2s^2 2p^6 3s^2 3p^6 3d^9 4s^2$$

The correct electron configurations are as follows:

$$\text{Cr } 1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$$
$$\text{Cu } 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$$

These arrangements give chromium a half-filled *d* sublevel and copper a filled *d* sublevel. Filled energy sublevels are more stable than partially filled sublevels. Some actual electron configurations differ from those assigned using the aufbau principle because although half-filled sublevels are not as stable as filled sublevels, they are more stable than other configurations. This tendency overcomes the small difference between the energies of the 3*d* and 4*s* sublevels in copper and chromium.

At higher principal quantum numbers, energy differences between some sublevels (such as 5*f* and 6*d*, for example) are even smaller than in the chromium and copper examples. As a result, there are other exceptions to the aufbau principle. Although it is worth knowing that exceptions to the aufbau principle occur, it is more important to understand the general rules for determining electron configurations in the many cases in which the aufbau principle applies.

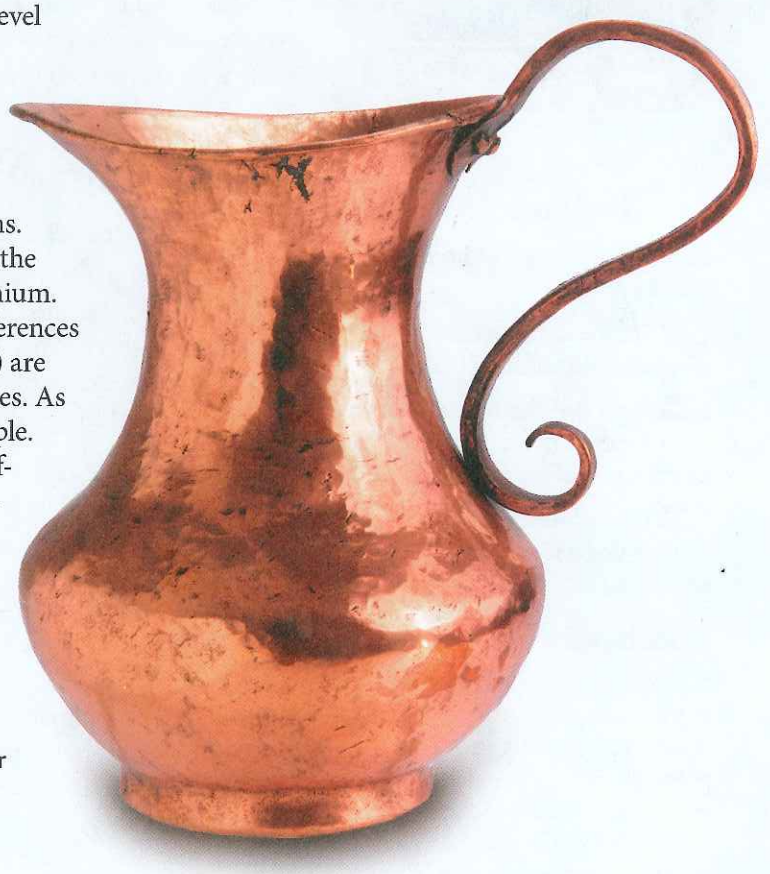


Figure 5.6 Copper
Copper is a shiny metal that can be molded into different shapes. The electron configuration of copper does not follow the aufbau principle.

5.2 LessonCheck

ONLINE PROBLEMS

10. List What are the three rules for writing the electron configurations of elements?

11. Sequence Use Figure 5.5 to arrange the following sublevels in order of decreasing energy: 2*p*, 4*s*, 3*s*, 3*d*, and 3*p*.

12. Explain Why do the actual electron configurations for some elements differ from those assigned using the aufbau principle?

13. Infer Why does one electron in a potassium atom go into the fourth energy level instead of squeezing into the third energy level along with the eight already there?

14. Apply Concepts The atomic number of arsenic is 33. What is the electron configuration of an arsenic atom?

5.3 Atomic Emission Spectra and the Quantum Mechanical Model

CHEMISTRY & YOU

Q: *What gives gas-filled lights their colors?* If you walk in the evening along a busy street lined with shops and theaters, you are likely to see lighted advertising signs. The signs are formed from glass tubes bent in various shapes. An electric current passing through the gas in each glass tube makes the gas glow with its own characteristic color.

Key Questions

What causes atomic emission spectra?

How did Einstein explain the photoelectric effect?

How are the frequencies of light emitted by an atom related to changes of electron energies?

How does quantum mechanics differ from classical mechanics?

Vocabulary

- amplitude
- wavelength
- frequency
- hertz
- electromagnetic radiation
- spectrum
- atomic emission spectrum
- Planck's constant
- photoelectric effect
- photon
- ground state
- Heisenberg uncertainty principle

Light and Atomic Emission Spectra

What causes atomic emission spectra?

The previous sections in this chapter introduced you to some ideas about how electrons in atoms are arranged in orbitals, each with a particular energy level. You also learned how to write electron configurations for atoms. You will now get a closer look into what led to the development of Schrödinger's equation and the quantum mechanical model of the atom.

The Nature of Light Rather curiously, the quantum mechanical model grew out of the study of light. Isaac Newton (1642–1727) tried to explain what was known about the behavior of light by assuming that light consists of particles. By the year 1900, however, there was enough experimental evidence to convince scientists that light consists of waves. Figure 5.7 illustrates some of the properties of waves. As shown, each complete wave cycle starts at zero on the y-axis, increases to its highest value, passes through zero to reach its lowest value, and returns to zero again. The **amplitude** of a wave is the wave's height from zero to the crest, as shown in Figure 5.7. The **wavelength,** represented by λ (the Greek letter lambda), is the distance between the crests. The **frequency,** represented by ν (the Greek letter nu), is the number of wave cycles to pass a given point per unit of time. The units of frequency are usually cycles per second. The SI unit of cycles per second is called the **hertz** (Hz). A hertz can also be expressed as a reciprocal second (s^{-1}).

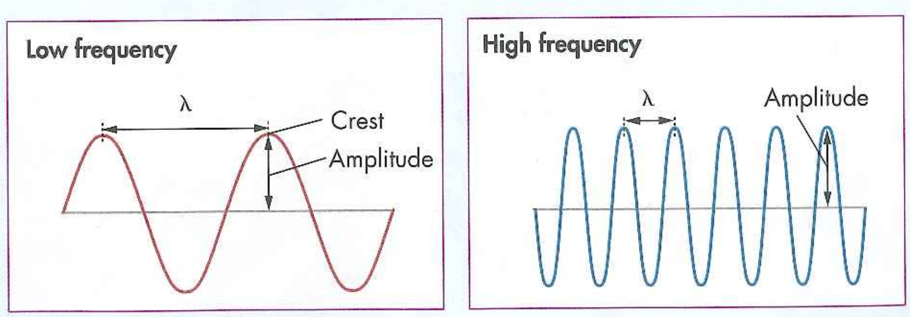


Figure 5.7 Light Waves
The frequency (ν) and wavelength (λ) of light waves are inversely related. As the wavelength decreases, the frequency increases.

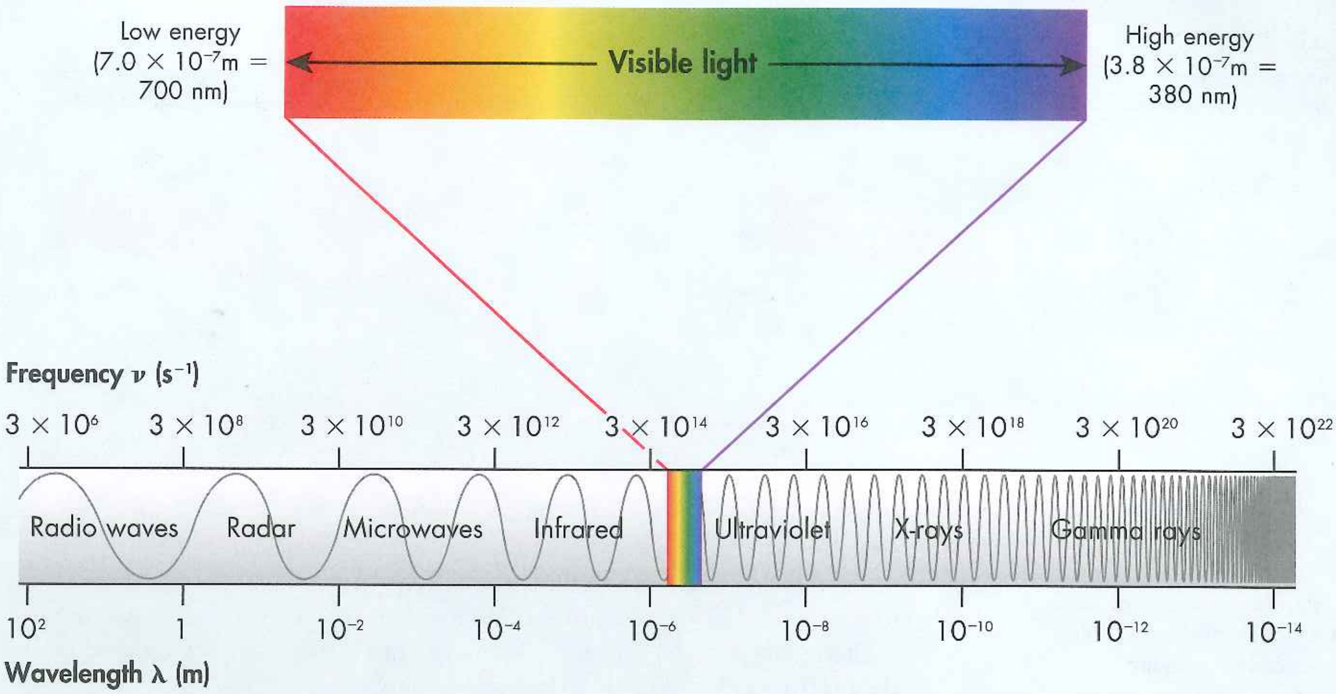


Figure 5.8 Electromagnetic Spectrum
The electromagnetic spectrum consists of radiation over a broad range of wavelengths. The visible light portion is very small. It is in the 10^{-7} m wavelength range and 10^{15} Hz (s^{-1}) frequency range.
Interpret Diagrams ***What types of nonvisible radiation have wavelengths close to those of red light? To those of blue light?***

The product of frequency and wavelength equals a constant (c), the speed of light.

$$c = \lambda\nu$$

The wavelength and frequency of light are inversely proportional to each other. As the wavelength of light increases, the frequency decreases.

According to the wave model, light consists of electromagnetic waves. **Electromagnetic radiation** includes radio waves, microwaves, infrared waves, visible light, ultraviolet waves, X-rays, and gamma rays. All electromagnetic waves travel in a vacuum at a speed of 2.998×10^8 m/s.

The sun and incandescent light bulbs emit white light, which consists of light with a continuous range of wavelengths and frequencies. As you can see from Figure 5.8, the wavelength and frequency of each color of light are characteristic of that color. When sunlight passes through a prism, the different wavelengths separate into a **spectrum** of colors. A rainbow is an example of this phenomenon. Each tiny droplet of water acts as a prism to produce a spectrum. Each color blends into the next in the order red, orange, yellow, green, blue, and violet. As can be seen in Figure 5.8, red light has the longest wavelength and the lowest frequency in the visible spectrum.

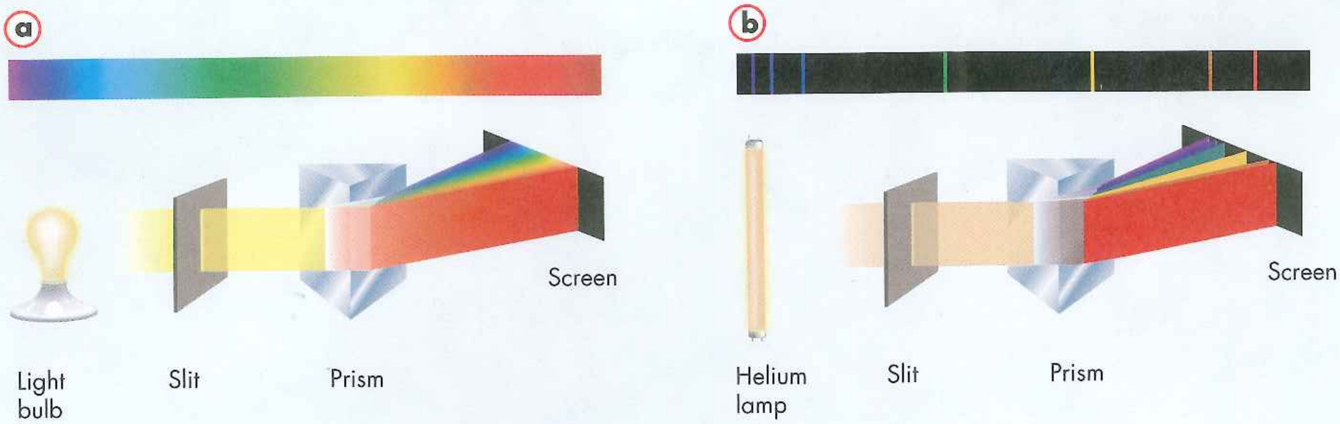


Figure 5.9 Comparing Spectra
A prism separates light into the colors it contains. **a.** White light produces a rainbow of colors. **b.** Light from a helium lamp produces discrete lines.
Identify *Which color of the rainbow has the highest frequency?*

Atomic Emission Spectra When an electric current is passed through a gaseous element, or through the vapor of a liquid or solid element, the electrons of the atoms of the gas or vapor are energized. This energy causes them to emit light. **When atoms absorb energy, their electrons move to higher energy levels. These electrons lose energy by emitting light when they return to lower energy levels.** The energy absorbed by an electron for it to move from its current energy level to a higher energy level is identical to the energy of the light emitted by the electron as it drops back to its original energy level. Figure 5.9a shows the visible spectrum of white light. Notice that all the wavelengths of visible light are blurred together as in a rainbow. However, when the light emitted by the energized electrons of a gaseous element is passed through a prism, as shown in Figure 5.9b, the spectrum consists of a limited number of narrow lines of light. The wavelengths of these spectral lines are characteristic of the element, and they make up the **atomic emission spectrum** of the element.

Each spectral line in an atomic emission spectrum of an element corresponds to exactly one wavelength of light emitted by the electrons of that element. Figure 5.9b shows the visible portion of the atomic emission spectrum of helium.

The atomic emission spectrum of each element is like a person's fingerprint. Just as no two people have the same fingerprints, no two elements have the same atomic emission spectrum. In the same way that fingerprints identify people, atomic emission spectra are useful for identifying elements. Figure 5.10 shows the characteristic colors emitted by sodium and by mercury. Much of the knowledge about the composition of the universe comes from studying the atomic emission spectra of the stars, which are hot glowing bodies of gases.

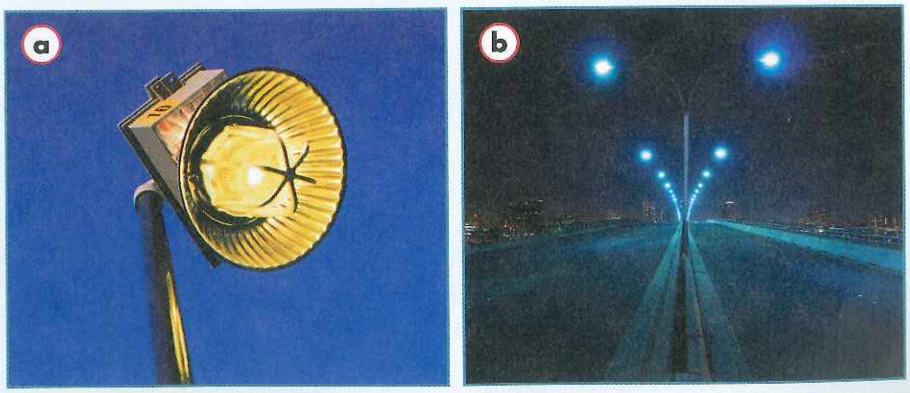


Figure 5.10
Atomic Emission Spectra
No two elements have the same atomic emission spectrum.
a. Sodium vapor lamps produce a yellow glow. **b.** Mercury vapor lamps produce a blue glow.

Sample Problem 5.2

Calculating the Wavelength of Light

Calculate the wavelength of the yellow light emitted by a sodium lamp if the frequency of the radiation is 5.09×10^{14} Hz (5.09×10^{14}/s).

1 Analyze **List the knowns and the unknown.** Use the equation $c = \lambda\nu$ to solve for the unknown wavelength.

KNOWNS

frequency (ν) = 5.09×10^{14}/s

$c = 2.998 \times 10^{8}$ m/s

UNKNOWN

wavelength (λ) = ? m

2 Calculate **Solve for the unknown.**

Write the expression that relates the frequency and wavelength of light.

$$c = \lambda\nu$$

Solve for λ by dividing both sides by ν:

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

Rearrange the equation to solve for λ.

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

Substitute the known values for ν and c into the equation and solve.

$$\lambda = \frac{2.998 \times 10^{8}\ \text{m}/\cancel{\text{s}}}{5.09 \times 10^{14}/\cancel{\text{s}}} = \boxed{5.89 \times 10^{-7}\ \text{m}}$$

3 Evaluate **Does the result make sense?** The magnitude of the frequency is much larger than the numerical value of the speed of light, so the answer should be much less than 1. The answer should have three significant figures.

15. What is the wavelength of radiation with a frequency of 1.50×10^{13} Hz? Does this radiation have a longer or shorter wavelength than red light?

16. What is the frequency of radiation with a wavelength of 5.00×10^{-8} m? In what region of the electromagnetic spectrum is this radiation?

In Problem 15, solve for wavelength.

In Problem 16, solve for frequency.

Quick Lab

Flame Tests

Purpose To determine the identity of the metal in an unknown solution based on its characteristic color in a flame

Materials

- 6 small test tubes
- sodium chloride ($NaCl$) solution
- calcium chloride ($CaCl_2$) solution
- lithium chloride ($LiCl$) solution
- copper(II) chloride ($CuCl_2$) solution
- potassium chloride (KCl) solution
- unknown solution
- 6 cotton swabs
- gas burner

Procedure

1. Make a two-column data table. Label the columns Metal and Flame Color. Enter the metal's name for each solution in the first column.

2. Label each of five test tubes with the name of a solution; label the sixth tube Unknown. Add 1 mL of each solution to the appropriately labeled test tube.

3. Dip one of the cotton ends of a cotton swab into the sodium chloride solution and then hold it briefly in the burner flame. Record the color of the flame. Do not leave the swab in the flame too long or the plastic will melt.

4. Repeat Step 3 for each of the remaining solutions using a new cotton swab each time.

5. Perform a flame test with the unknown solution. Note the color of the flame.

Analyze and Conclude

1. Identify What is the metal in the unknown?

2. Draw Conclusions Each solution produces a unique color. Would you expect this result based on the modern view of the atom? Explain.

3. Analyze Data Some commercially available fireplace logs burn with a red and/or green flame. What elements could be responsible for these colored flames?

4. Predict Aerial fireworks contain gunpowder and chemicals that produce colors. What element would you include to produce crimson red? Yellow?

The Quantum Concept and Photons

How did Einstein explain the photoelectric effect?

According to the laws of classical physics, the atomic emission spectrum of an element should be continuous. Thus, classical physics does not explain the emission spectra of atoms, which consist of lines.

The Quantization of Energy Recall the iron scroll in Figure 5.1 that changed color when heated. In 1900, the German physicist Max Planck (1858–1947) was trying to describe why such a body first appears black, then red, then yellow, and then white as its temperature increases. Planck found that he could explain the color changes if he assumed that the energy of a body changes only in small discrete units, or quanta. Planck showed mathematically that the amount of radiant energy (E) of a single quantum absorbed or emitted by a body is proportional to the frequency of radiation (ν).

$$E \propto \nu \text{ or } E = h\nu$$

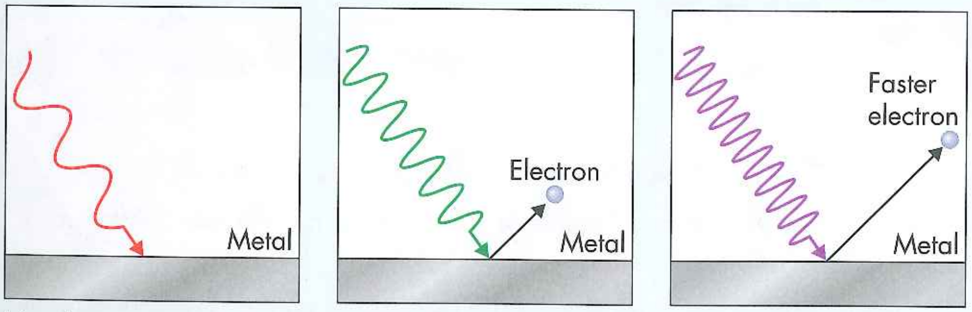
The constant (h), which has a value of 6.626×10^{-34} J·s, (J is the joule, the SI unit of energy) is called **Planck's constant.** The energy of a quantum equals $h\nu$. A small energy change involves the emission or absorption of low-frequency radiation. A large energy change involves the emission or absorption of high-frequency radiation.

The Photoelectric Effect A few years after Planck presented his theory on the quantization of energy, scientists began to use it to explain many experimental observations that could not be explained by classical physics. In 1905, Albert Einstein (1879–1955), then a patent examiner in Bern, Switzerland, used Planck's quantum theory to explain the photoelectric effect, which is illustrated in Figure 5.11. In the **photoelectric effect,** electrons are ejected when light shines on a metal. Not just any frequency of light will cause the photoelectric effect. For example, red light will not cause potassium to eject electrons, no matter how intense the light. Yet a very weak yellow light shining on potassium begins the effect.

The photoelectric effect could not be explained by classical physics. Although classical physics correctly described light as a form of energy, it assumed that under weak light of any wavelength, an electron in a metal should eventually collect enough energy to be ejected. The photoelectric effect presented a serious problem for the classical wave theory of light.

To explain the photoelectric effect, Einstein proposed that light could be described as quanta of energy that behave as if they were particles. These light quanta are called **photons.** The energy of photons is quantized according to the equation $E = h\nu$. Einstein recognized that there is a threshold value of energy below which the photoelectric effect does not occur. According to $E = h\nu$, all the photons in a beam of monochromatic light (light of only one frequency) have the same energy. If the frequency, and therefore the energy, of the photons is too low, then no electrons will be ejected. It does not matter whether a single photon or a steady stream of low-energy photons strikes an electron in the metal. Only if the frequency of light is above the threshold frequency will the photoelectric effect occur.

Einstein's theory that light behaves as a stream of particles explains the photoelectric effect and many other observations. However, light also behaves as waves in other situations. Therefore, we must consider that light possesses both wavelike and particle-like properties.



No electrons are ejected because the frequency of the light is below the threshold frequency.

If the light is at or above the threshold frequency, electrons are ejected.

If the frequency is increased, the ejected electrons will travel faster.

Figure 5.11
Photoelectric Effect
Einstein explained the photoelectric effect by proposing that light behaves as particles.
Predict ***What will happen if ultraviolet light shines on the metal?***

Sample Problem 5.3

Calculating the Energy of a Photon

What is the energy of a photon of microwave radiation with a frequency of 3.20×10^{11}/s?

KNOWNS

frequency (ν) = 3.20×10^{11}/s

$h = 6.626 \times 10^{-34}$ J·s

UNKNOWN

energy (E) = ? J

1 Analyze **List the knowns and the unknown.** Use the equation $E = h\nu$ to calculate the energy of the photon.

2 Calculate **Solve for the unknown.**

Write the expression that relates the energy of a photon of radiation and the frequency of the radiation.

$$E = h\nu$$

Substitute the known values for ν and h into the equation and solve.

$$E = (6.626 \times 10^{-34}\ \text{J}\cdot\cancel{\text{s}}) \times (3.20 \times 10^{11}/\cancel{\text{s}})$$

$$= 2.12 \times 10^{-22}\ \text{J}$$

3 Evaluate **Does the result make sense?** Individual photons have very small energies, so the answer seems reasonable.

17. Calculate the energy of a quantum of radiant energy with a frequency of 5.00×10^{11}/s.

18. The threshold photoelectric effect in tungsten is produced by light of a wavelength 260 nm. Give the energy of a photon of this light in joules.

An Explanation of Atomic Spectra

How are the frequencies of light emitted by an atom related to changes of electron energies?

Atomic emission spectra were known before Bohr proposed his model of the hydrogen atom. Bohr applied quantum theory to electron energy levels in atoms to explain the atomic emission spectrum of hydrogen. Bohr's model not only explained why the atomic emission spectrum of hydrogen consists of specific frequencies of light, but it also predicted specific values of these frequencies that agreed with the experimental results.

In the Bohr model, the lone electron in the hydrogen atom can have only certain specific energies. When the electron has its lowest possible energy, the atom is in its **ground state.** In the ground state, the principal quantum number (n) is 1. Excitation of the electron by absorbing energy raises the atom to an excited state with $n = 2, 3, 4, 5$, or 6, and so forth. A quantum of energy in the form of light is emitted when the electron drops back to a lower energy level. The emission occurs in a single step, called an electronic transition. Bohr already knew that this quantum of energy E is related to the frequency ν of the emitted light by the equation $E = h\nu$. **The light emitted by an electron moving from a higher to a lower energy level has a frequency directly proportional to the energy change of the electron.** Therefore, each transition produces a line of a specific frequency in the spectrum.

Figure 5.12 shows the three groups of lines in the emission spectrum of hydrogen atoms. The lines at the ultraviolet end of the hydrogen spectrum are the Lyman series. These lines are due to the transitions of electrons from higher energy levels to the lowest energy level, $n = 1$. The lines in the visible spectrum are the Balmer series. These lines result from transitions from higher energy levels to $n = 2$. These transitions generally involve a smaller change in electron energy than transitions to $n = 1$. Transitions to $n = 3$ from higher energy levels produce the Paschen series. The energy changes of the electron are generally smaller still. The lines are in the infrared range. Spectral lines for the transitions from higher energy levels to $n = 4$ and $n = 5$ also exist. Note that the spectral lines in each group become more closely spaced at increased values of n because the energy levels become closer together. There is an upper limit to the frequency of emitted light for each set of lines because an electron with enough energy completely escapes the atom.

Bohr's model explained the atomic emission spectrum of hydrogen but not the emission spectra of atoms with more than one electron. Also, it did not help in understanding how atoms bond to form molecules. Eventually the quantum mechanical model displaced the Bohr model of the atom.

CHEMISTRY & YOU

Q: *The glass tubes in lighted signs contain helium, neon, argon, krypton, or xenon gas, or a mixture of these gases. Why do the colors of the light depend on the gases that are used?*

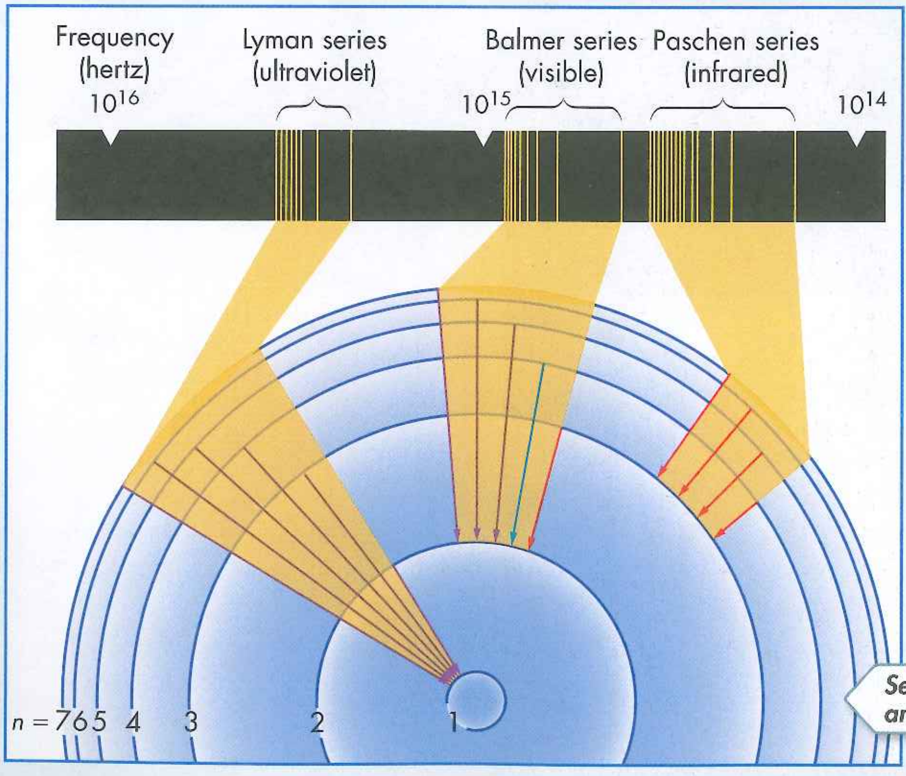
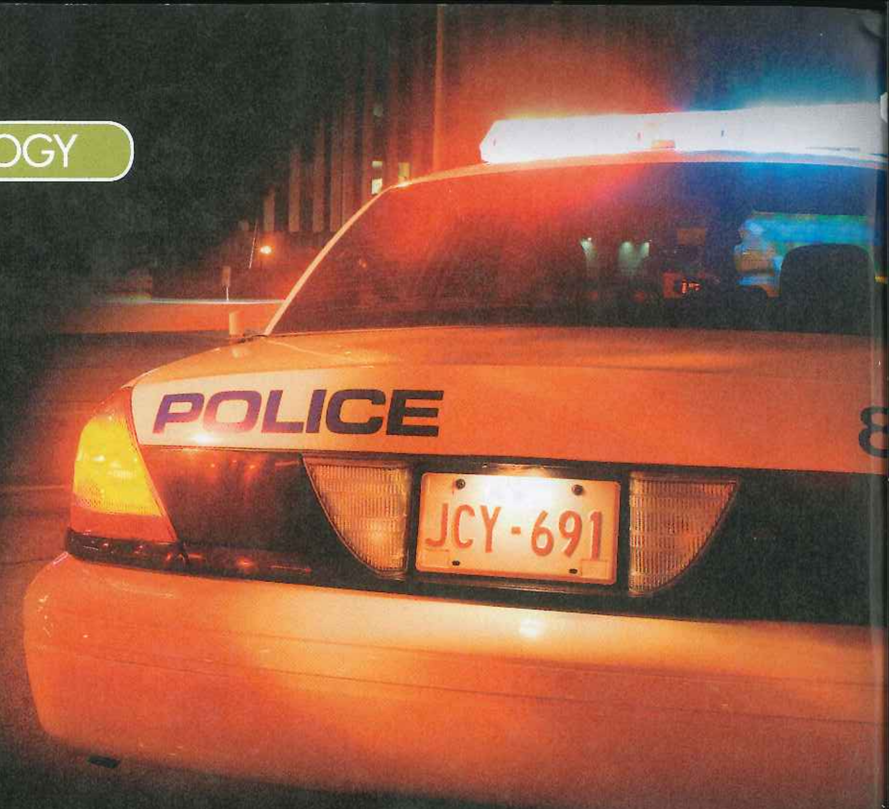


Figure 5.12 The Hydrogen Spectrum The three groups of lines in the hydrogen spectrum correspond to the transitions of electrons from higher energy levels to lower energy levels.
Interpret Diagrams ***Which of the following transitions produces the spectral line having the longest wavelength (lowest frequency): n = 2 to n = 1, n = 3 to n = 2, or n = 4 to n = 3?***

See the hydrogen spectrum animated online. KINETIC ART

CHEMISTRY & YOU: TECHNOLOGY

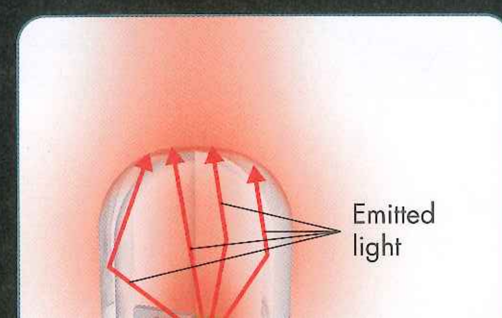
Light Emitting Diodes



Although they are small, you may have seen light emitting diodes, or LEDs, several times today. These tiny light bulbs may form the numbers on your digital clock, light up your watch, or illuminate the traffic light you stopped at on your way to school. You may have even watched television on a giant screen made out of of LEDs.

Light from a typical incandescent bulb is generated when the filament inside the bulb is heated. Light from an LED is generated in a different way. A diode is made out of two materials with different properties. The electrons in one of the materials are at a higher energy level than the electrons in the other material. When a voltage, supplied by a battery or other power supply, is applied to the diode, electrons flow across the boundary between the two materials. The electrons at the higher energy level flow into the other material, fall to a lower energy level, and emit light.

LEDs last much longer than incandescent light bulbs because there is no filament to burn out. Also, a large amount of the energy that is used to produce light in an incandescent bulb is wasted as heat. However, LEDs produce very little heat, so they use less energy and cost less to operate than incandescent lights.



Take It Further

1. **Apply Concepts** LEDs that produce infrared light can be used to transmit information from remote

Quantum Mechanics

How does quantum mechanics differ from classical mechanics?

In 1924, Louis de Broglie (1892–1987), a French graduate student, asked an important question: Given that light behaves as waves and particles, can particles of matter behave as waves? De Broglie referred to the wavelike behavior of particles as matter waves. His reasoning led him to a mathematical expression for the wavelength of a moving particle.

The Wavelike Nature of Matter The proposal that matter moves in a wavelike way would not have been accepted unless experiments confirmed its validity. Only three years later, experiments by Clinton Davisson and Lester Germer at Bell Labs in New Jersey did just that. The two scientists had been studying the bombardment of metals with beams of electrons. They noticed that the electrons reflected from the metal surface produced curious patterns. The patterns were like those obtained when X-rays (which are electromagnetic waves) reflect from metal surfaces. The electrons, which were believed to be particles, were reflected as if they were waves! De Broglie was awarded the Nobel prize for his work on the wave nature of matter. Davisson also received the Nobel prize for his experiments demonstrating the wave nature of electrons.

Today, the wavelike properties of beams of electrons are useful in viewing objects that cannot be viewed with an optical microscope. The electrons in an electron microscope have much smaller wavelengths than visible light. These smaller wavelengths allow a much clearer enlarged image of a very small object, such as the pollen grain in Figure 5.13, than is possible with an ordinary microscope.

De Broglie's equation predicts that all moving objects have wavelike behavior. Why are you unable to observe the effects of this wavelike motion for ordinary objects like baseballs and trains? The answer is that the mass of the object must be very small in order for its wavelength to be large enough to observe. For example, a 50-gram golf ball traveling at 40 m/s (about 90 mi/h) has a wavelength of only 3×10^{-34} m, which is much too small to detect experimentally. On the other hand, an electron has a mass of only 9.11×10^{-28} g. If it were moving at a velocity of 40 m/s, it would have a wavelength of 2×10^{-5} m, which is comparable to infrared radiation and is readily measured.

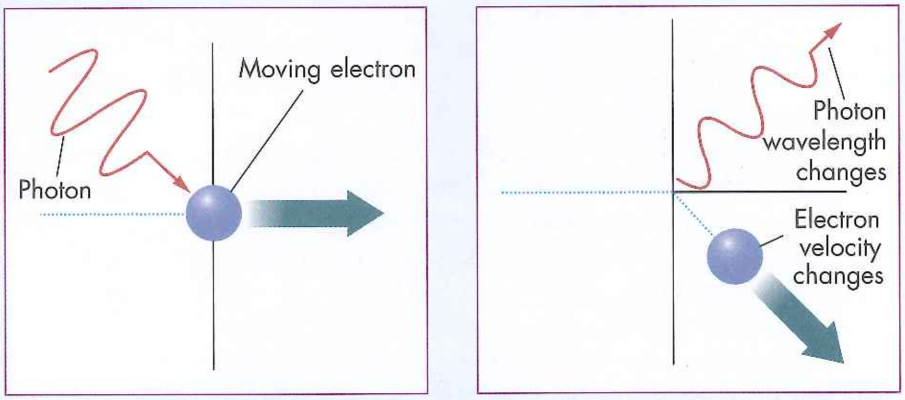
De Broglie's prediction that matter exhibits both wave and particle properties set the stage for a new way of describing the motions of subatomic particles and atoms. The newer theory is called quantum mechanics; the older theory is called classical mechanics. **Classical mechanics adequately describes the motions of bodies much larger than atoms, while quantum mechanics describes the motions of**

The Heisenberg Uncertainty Principle German physicist Werner Heisenberg examined another feature of quantum mechanics. The **Heisenberg uncertainty principle** states that it is impossible to know both the velocity and the position of a particle at the same time. This limitation is critical when dealing with small particles such as electrons, but it does not matter for ordinary-sized objects such as cars or airplanes.

Consider how you determine the location of an object. To locate a set of keys in a dark room you can use a flashlight. You see the keys when the light bounces off them and strikes your eyes. To locate an electron, you might strike it with a photon of light, as shown in Figure 5.14. However, the electron has such a small mass that striking it with a photon affects its motion in a way that cannot be predicted accurately. The very act of measuring the position of the electron changes its velocity, making its velocity uncertain.

The discovery of matter waves led the way for Schrödinger's quantum mechanical description of electrons in atoms. Schrödinger's theory leads to the concept of electron orbitals and includes the uncertainty principle.

Figure 5.14 Heisenberg Uncertainty Principle
According to the Heisenberg uncertainty principle, it is impossible to know exactly both the velocity and the position of a particle at the same time.



Before collision A photon strikes an electron during an attempt to observe the electron's position.

After collision The impact changes the electron's velocity, making it uncertain.

5.3 LessonCheck

19. **Describe** What is the origin of the atomic emission spectrum of an element?

20. **Review** What was Einstein's explanation for the photoelectric effect?

21. **Explain** How is the change in electron energy related to the frequency of light emitted in electronic transitions?

22. **Explain** How does quantum mechanics differ from classical mechanics?

23. **Sequence** Arrange the following in order of decreasing wavelength.
- **a.** infrared radiation from a heat lamp
- **b.** dental X-rays
- **c.** signal from a shortwave radio station

24. **Calculate** A hydrogen lamp emits several lines in the visible region of the spectrum. One of these lines has a wavelength of 6.56×10^{-5} cm. What is the frequency of this radiation?

25. **Calculate** What is the energy of a photon of blue light with a wavelength of 460 nm?

BIGIDEA
ELECTRONS AND THE STRUCTURE OF ATOMS

26. When a strontium compound is heated in a flame, red light is produced. When a barium compound is heated in a flame, yellow-green light is produced. Explain why these colors are emitted.

Small-Scale Lab

Atomic Emission Spectra

Purpose

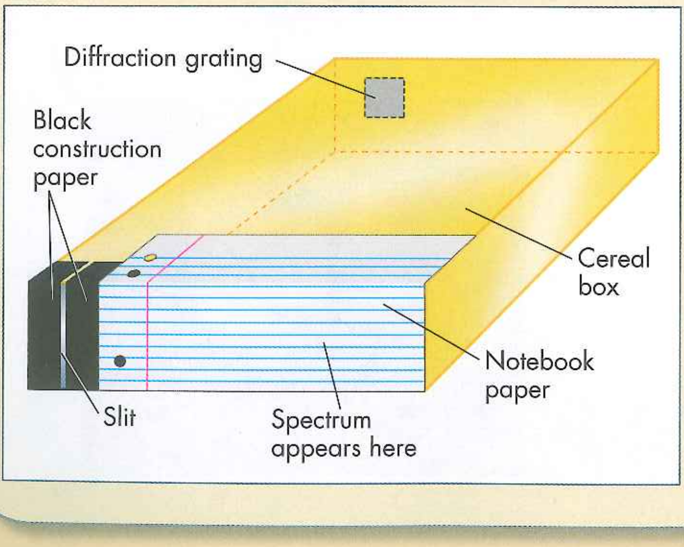
To build a spectroscope and use it to measure the wavelengths, frequencies, and energies of atomic emission lines

Materials

- **black construction paper**
- **tape**
- **cereal box**
- **scissors**
- **white notebook paper**
- **diffraction grating**
- **ruler**

Procedure

Tape together two 2.0 cm × 10 cm strips of black construction paper so that they are parallel and form a narrow slit about 2 mm wide. Remove the top of a cereal box and tape the construction paper slit as shown. Cover the rest of the opening with white notebook paper. Cut a square hole (approximately 2 cm per side) and tape a diffraction grating over the hole as shown. Point the spectroscope toward a fluorescent light. Tape up any light leaks. Your lab partner should mark the exact positions of all the colored emission lines you see on the notebook paper. Measure the distances between the violet line and the other lines you have marked.



Analyze and Conclude

1. Observe List the number of distinct lines that you see as well as their colors.

2. Measure Each line you see has a wavelength. The prominent violet line has a wavelength of 436 nm and the prominent green line has a wavelength of 546 nm. How many mm apart are these lines on the paper? By how many nm do their wavelengths differ? How many nanometers of wavelength are represented by each millimeter you measured?

3. Calculate Using the nm/mm value you calculated and the mm distance you measured for each line from the violet reference line, calculate the wavelengths of all the other lines you see.

4. Calculate Use the wavelength value of each line to calculate its frequency given that $\nu = c/\lambda$ where $c = 2.998 \times 10^{17}$ nm/s (2.998×10^{8} m/s).

5. Calculate The energy E of a quantum of light an atom emits is related to its frequency ν by $E = h\nu$. Use the frequency value for each line and $h = 6.626 \times 10^{-34}$ J·s to calculate its corresponding energy.

You're the Chemist

1. Design an Experiment Design and carry out an experiment to measure the longest and shortest wavelengths you can see in daylight. Use your spectroscope to observe light from daylight reflected off a white piece of paper. **CAUTION** *Do not look directly at the sun!* Describe the differences in daylight and fluorescent light.

2. Design an Experiment Design and carry out an experiment to determine the effect of colored filters on the spectrum of fluorescent light or daylight. For each filter, tell which colors are transmitted and which are absorbed.

3. Analyze Data Use your spectroscope to observe various atomic emission discharge tubes provided by your teacher. Note and record the lines you see and measure their wavelengths.

5 Study Guide

BIGIDEA ELECTRONS AND THE STRUCTURE OF ATOMS

The quantum mechanical model of the atom comes from the solutions to the Schrödinger equation. Solutions to the Schrödinger equation give the energies an electron can have and the atomic orbitals, which describe the regions of space where an electron may be found. Electrons can absorb energy to move from one energy level to a higher energy level. When an electron moves from a higher energy level back down to a lower energy level, light is emitted.

5.1 Revising the Atomic Model

Bohr proposed that an electron is found only in specific circular paths, or orbits, around the nucleus.

The quantum mechanical model determines the allowed energies an electron can have and how likely it is to find the electron in various locations around the nucleus of an atom.

Each energy sublevel corresponds to one or more orbitals of different shapes. The orbitals describe where an electron is likely to be found.

- energy level (129)
- quantum (129)
- quantum mechanical model (130)
- atomic orbital (131)

5.2 Electron Arrangement in Atoms

Three rules—the aufbau principle, the Pauli exclusion principle, and Hund's rule—tell you how to find the electron configurations of atoms.

- electron configuration (134)
- aufbau principle (134)
- Pauli exclusion principle (134)
- spin (134)
- Hund's rule (134)

5.3 Atomic Emission Spectra and the Quantum Mechanical Model

When atoms absorb energy, their electrons move to higher energy levels. These electrons lose energy by emitting light when they return to lower energy levels.

To explain the photoelectric effect, Einstein proposed that light could be described as quanta of energy that behave as if they were particles.

The light emitted by an electron moving from a higher to a lower energy level has a frequency directly proportional to the energy change of the electron.

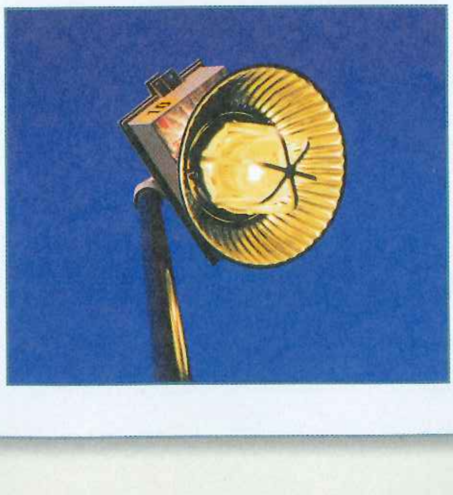
Classical mechanics adequately describes the motions of bodies much larger than atoms, while quantum mechanics describes the motions of subatomic particles and atoms as waves.

- amplitude (138)
- wavelength (138)
- frequency (138)
- hertz (138)
- electromagnetic radiation (139)
- spectrum (139)
- atomic emission spectrum (140)
- Planck's constant (143)
- photoelectric effect (143)
- photon (143)
- ground state (145)
- Heisenberg uncertainty principle (148)

Key Equations

$$c = \lambda \nu$$

$$E = h\nu$$



Math Tune-Up: Atomic Emission Spectra and Photons

MATH +/– TUTOR

Problem

Calculate the wavelength of radiation with a frequency of 8.43×10^9 Hz (8.43×10^9/s). In what region of the electromagnetic spectrum is this radiation?

What is the energy of a photon of X-ray radiation with a frequency of 7.49×10^{18}/s?

1 Analyze

Knowns:
$\nu = 8.43 \times 10^9/\text{s}$
$c = 2.998 \times 10^8 \text{ m/s}$

Unknown:
$\lambda = ?\ \text{m}$

Use the equation that relates the frequency and wavelength of light:
$c = \lambda\nu$

Knowns:
$\nu = 7.49 \times 10^{18}/\text{s}$
$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$

Unknown:
$E = ?\ \text{J}$

Use the equation that relates the energy of a photon of radiation and the frequency of the radiation:
$E = h\nu$

2 Calculate

Solve for λ and calculate.

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

$$\lambda = \frac{2.998 \times 10^8 \text{ m}/\cancel{\text{s}}}{8.43 \times 10^9/\cancel{\text{s}}}$$

$$\lambda = 3.56 \times 10^{-2} \text{ m}$$

The radiation is in the radar region of the electromagnetic spectrum.

Substitute the known values for ν and h into the equation and calculate.

$$E = (6.626 \times 10^{-34} \text{ J}\cdot\cancel{\text{s}}) \times (7.49 \times 10^{18}/\cancel{\text{s}})$$

$$E = 4.96 \times 10^{-15} \text{ J}$$

If you are given the wavelength of the radiation, first calculate frequency using $c = \lambda\nu$, and then use $E = h\nu$ to calculate energy.

3 Evaluate

The magnitude of the frequency of the radiation is larger than the value for the speed of light, so the answer should be less than 1.

Hint: Review Sample Problem 5.2 if you have trouble with converting between wavelength and frequency.

Individual photons have very small energies, so the answer is reasonable.

5 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

5.1 Revising the Atomic Model

27. Why was Rutherford's model of the atom known as the planetary model?

★**28.** What did Bohr assume about the motion of electrons?

29. Describe Rutherford's model of the atom and compare it with the model proposed by his student Niels Bohr.

★**30.** What is the significance of the boundary of an electron cloud?

31. What is an atomic orbital?

32. Sketch 1*s*, 2*s*, and 2*p* orbitals using the same scale for each.

★**33.** How many orbitals are in the 2*p* sublevel?

★**34.** How many sublevels are contained in each of these principal energy levels?

a. $n = 1$
b. $n = 2$
c. $n = 3$
d. $n = 4$

5.2 Electron Arrangement in Atoms

★**35.** What are the three rules that govern the filling of atomic orbitals by electrons?

★**36.** Arrange the following sublevels in order of increasing energy:
3*d*, 2*s*, 4*s*, 3*p*.

★**37.** Which of these orbital designations are invalid?

a. 4*s*
b. 2*d*
c. 3*f*
d. 3*p*

38. What is the maximum number of electrons that can go into each of the following sublevels?

a. 2*s*
b. 4*s*
c. 4*p*
d. 4*f*
e. 3*p*
f. 3*d*
g. 5*s*
h. 5*p*

★**39.** What is meant by $3p^3$?

40. Write electron configurations for the elements that are identified by these atomic numbers:

a. 7
b. 9
c. 12
d. 36

41. Give electron configurations for atoms of these elements:

a. Na
b. K
c. I
d. Ne

★**42.** How many electrons are in the highest occupied energy level of these atoms?

a. barium
b. aluminum
c. sodium
d. oxygen

43. How many electrons are in the second energy level of an atom of each element?

a. chlorine
b. phosphorus
c. potassium

★**44.** Write electron configurations for atoms of these elements:

a. selenium
b. titanium
c. vanadium
d. calcium

5.3 Atomic Emission Spectra and the Quantum Mechanical Model

45. Use a diagram to illustrate each term for a wave.

a. wavelength
b. amplitude
c. cycle

46. What is meant by the frequency of a wave? What are the units of frequency? Describe the relationship between frequency and wavelength.

★**47.** Consider the following regions of the electromagnetic spectrum: (i) ultraviolet, (ii) X-ray, (iii) visible, (iv) infrared, (v) radio wave, (vi) microwave.

a. Use Figure 5.8 to arrange them in order of decreasing wavelength.
b. How does this order differ from that of decreasing frequency?

48. List the colors of the visible spectrum in order of increasing wavelength.

49. How did Planck influence the development of modern atomic theory?

★**50.** Explain the difference between a photon and a quantum.

★**51.** What has more energy, a photon of infrared light or a photon of ultraviolet light?

*52. What is the energy of a photon of green light with a frequency of 5.80×10^{14}/s?

*53. Explain the difference between the energy lost or gained by an atom according to the laws of classical physics and according to the quantum model of the atom.

*54. What happens when a hydrogen atom absorbs a quantum of energy?

55. The transition of electrons from higher energy levels to the $n = 2$ energy level results in the emission of light from hydrogen atoms. In what part of the spectrum is the emitted light, and what is the name given to this transition series?

Understand Concepts

56. Give the symbol for the atom that corresponds to each electron configuration.
   a. $1s^22s^22p^63s^23p^6$
   b. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^75s^1$
   c. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^75s^25p^65d^16s^2$

*57. Write the electron configuration for an arsenic atom. Calculate the total number of electrons in each energy level and state which energy levels are not full.

58. How many paired electrons are there in an atom of each element?
   a. helium
   b. sodium
   c. boron
   d. oxygen

*59. An atom of an element has two electrons in the first energy level and five electrons in the second energy level. Write the electron configuration for this atom and name the element. How many unpaired electrons does an atom of this element have?

*60. Give the symbols and names of the elements that correspond to these configurations of an atom.
   a. $1s^22s^22p^63s^1$
   b. $1s^22s^22p^3$
   c. $1s^22s^22p^63s^23p^2$
   d. $1s^22s^22p^4$
   e. $1s^22s^22p^63s^23p^64s^1$
   f. $1s^22s^22p^63s^23p^63d^24s^2$

61. What is the maximum number of electrons that can be found in any orbital of an atom?

62. Suppose your favorite AM radio station broadcasts at a frequency of 1150 kHz. What is the wavelength, in centimeters, of the radiation from the station?

*63. A mercury lamp, such as the one below, emits radiation with a wavelength of 4.36×10^{-7} m.

   a. What is the wavelength of this radiation in centimeters?
   b. In what region of the electromagnetic spectrum is this radiation?
   c. Calculate the frequency of this radiation.

64. Sodium vapor lamps are used to illuminate streets and highways. The very bright light emitted by these lamps is actually due to two closely spaced emission lines in the visible region of the electromagnetic spectrum. One of these lines has a wavelength of 5.890×10^{-7} m, and the other line has a wavelength of 5.896×10^{-7} m.
   a. What are the wavelengths of these radiations in centimeters?
   b. Calculate the frequencies of these radiations.
   c. In what region of the visible spectrum do these lines appear?

*65. What will happen if the following occur?
   a. Monochromatic light shining on cesium metal is just above the threshold frequency.
   b. The intensity of the light increases, but the frequency remains the same.
   c. Monochromatic light of a shorter wavelength is used.

*66. Calculate the energy of a photon of red light with a wavelength of 6.45×10^{-5} cm. Compare your answer with the answer to Question 52. Is red light of higher or lower energy than green light?

67. State the Heisenberg uncertainty principle.

68. Describe how the wavelength of a wave changes if the frequency of the wave is multiplied by 1.5.

*69. Indicate whether each of the following electron transitions emits energy or requires the absorption of energy.

a. $3p$ to $3s$
b. $3p$ to $4p$
c. $2s$ to $2p$
d. $1s$ to $2s$

*70. White light is viewed in a spectroscope after passing through sodium vapor too cool to emit light. The spectrum is continuous except for a dark line at 589 nm. How can you explain this observation? (*Hint:* Recall from Sample Problem 5.2 that the atomic emission spectrum of sodium exhibits a strong yellow line at 589 nm.)

71. You use a microwave oven to heat your dinner. The frequency of the radiation is $2.37 \times 10^9\ s^{-1}$. What is the energy of one photon of this radiation?

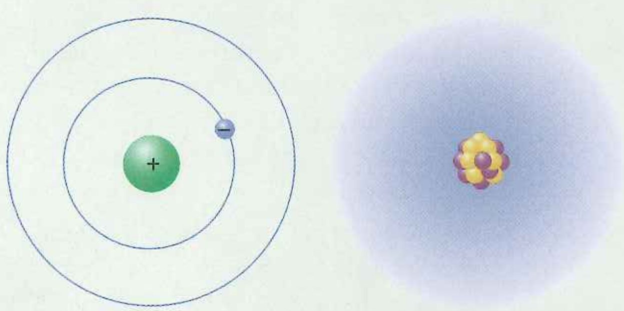
*72. Calculate the following energies:

a. One photon of infrared radiation, if $\lambda = 1.2 \times 10^{-4}$ m.
b. One photon of visible radiation, if $\lambda = 5.1 \times 10^{-7}$ m.
c. One photon of ultraviolet radiation, if $\lambda = 1.4 \times 10^{-8}$ m.

What do the answers indicate about the relationship between the energy of light and its wavelength?

Think Critically

*73. **Compare** Explain the difference between an orbit in the Bohr model and an orbital in the quantum mechanical model of the atom.



74. **Apply Concepts** Identify the elements whose electrically neutral atoms have the following electron configurations.

a. $1s^2 2s^2 2p^5$
b. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$
c. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$

*75. **Predict** Traditional cooking methods make use of infrared radiation (heat). Microwave radiation cooks food faster. Could radio waves be used for cooking? Explain.

76. **Draw Conclusions** Think about the currently accepted models of the atom and of light. In what ways do these models seem strange to you? Why are these models not exact or definite?

77. **Evaluate and Revise** Orbital diagrams for the ground states of two elements are shown below. Each diagram shows something that is incorrect. Identify the error in each diagram and then draw the correct diagram.

a. Nitrogen

1s	2s	2p		
↑↓	↑↓	↑↓	↑	

b. Magnesium

1s	2s	2p			3s
↑↓	↑↓	↑↓	↑↓	↑↓	

*78. **Infer** Picture two hydrogen atoms. The electron in the first hydrogen atom is in the $n = 1$ level. The electron in the second atom is in the $n = 4$ level.

a. Which atom has the ground state electron configuration?
b. Which atom can emit electromagnetic radiation?
c. In which atom is the electron in a larger orbital?
d. Which atom has the lower energy?

*79. **Infer** Which of the following is the ground state of an atom? Which is its excited state? Which is an impossible electron configuration? Identify the element and briefly explain your choices.

a. $1s^2 2s^2 2p^6 3s^2 3p^6 5p^1$
b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
c. $1s^2 2s^2 2p^6 3s^2 3p^7$

80. **Relate Cause and Effect** Why do electrons occupy equal energy orbitals singly before beginning to pair up?

Enrichment

*81. **Graph** The energy of a photon is related to its frequency and its wavelength.

Energy of photon (J)	Frequency (s^{-1})	Wavelength (cm)
3.45×10^{-21}	ν_1 ________	5.77×10^{-3}
2.92×10^{-20}	ν_2 ________	6.82×10^{-4}
6.29×10^{-20}	ν_3 ________	3.16×10^{-4}
1.13×10^{-19}	ν_4 ________	1.76×10^{-4}
1.46×10^{-19}	ν_5 ________	1.36×10^{-4}
3.11×10^{-19}	ν_6 ________	6.38×10^{-5}

a. Complete the table above.
b. Plot the energy of the photon (y-axis) versus the frequency (x-axis).
c. Determine the slope of the line.
d. What is the significance of this slope?

82. **Calculate** The average distance between Earth and Mars is about 2.08×10^8 km. How long would it take to transmit television pictures from Mars to Earth?

*83. **Calculate** Bohr's atomic theory can be used to calculate the energy required to remove an electron from an orbit of a hydrogen atom or an ion (an atom that has lost or gained electrons) containing only one electron. This number is the ionization energy for that atom or ion. The formula for determining the ionization energy (E) is

$$E = Z^2 \times \frac{k}{n^2}$$

where Z is the atomic number, k is 2.18×10^{-18} J, and n is the energy level. What is the energy required to eject an electron from a hydrogen atom when the electron is in the ground state ($n = 1$)? In the second energy level? How much energy is required to eject a ground state electron from the species Li^{2+} (a lithium atom that has lost two electrons)?

84. **Draw Conclusions** In a photoelectric experiment, a student shines light on the surface of a metal. The frequency of the light is greater than the threshold frequency of the metal. The student observes that after a long time, the maximum energy of the ejected electrons begins to decrease. Explain this observation.

Write About Science

*85. **Explain** Write a brief description of how trying to place two bar magnets pointing in the same direction alongside each other is like trying to place two electrons into the same orbital.

86. **Connect to the BIGIDEA** The late 1800s and early 1900s were significant times for the rapid development of chemistry. Bohr improved on Rutherford's model of the atom, then Schrödinger developed the quantum mechanical model of the atom. Explain why a model of the atom is crucial to understanding chemistry and in explaining the behavior of matter.

CHEMYSTERY

Now You See It... Now You Don't

Liam eventually realized that his star stickers would always stop glowing after a period of time. He discovered that he could "recharge" the stickers by turning on the lights. After he turned off the lights, the stars would glow again. However, after a few hours, the stars would eventually stop glowing.

Glow-in-the-dark objects contain compounds that react with light. When these objects are exposed to light, the electrons in the compounds absorb energy and become excited. As the electrons drop back down to a lower energy level, they emit light. This process, called phosphorescence, occurs more slowly in the compounds contained in glow-in-the-dark objects than in other compounds.

87. **Infer** Do Liam's glow-in-the-dark stars glow when the lights are on? Explain.

*88. **Connect to the BIGIDEA** Light emitted from an incandescent light bulb is in the visible region of the electromagnetic spectrum (300 nm to 700 nm). What does this information tell you about the energy of the photons absorbed by the electrons in glow-in-the dark objects?

Cumulative Review

*89. Classify each of the following as homogeneous or heterogeneous:
  a. a page of this textbook
  b. a banana split
  c. the water in bottled water

90. Hamburger undergoes a chemical change when cooked on a grill. All chemical changes are subject to the law of conservation of mass. Yet, a cooked hamburger will weigh less than the uncooked meat patty. Explain.

*91. Homogeneous mixtures and compounds are both composed of two or more elements. How do you distinguish between a homogeneous mixture and a compound?

92. The photo shows a magnified view of a piece of granite. Is granite a substance or a mixture?

*93. The diameter of a carbon atom is 77 pm. Express this measurement in μm.

94. A silver bar has a mass of 368 g. What is the volume, in cm^3, of the bar? The density of silver is 19.5 g/cm^3.

*95. Which has more mass, a 28.0-cm^3 piece of lead or a 16.0-cm^3 piece of gold? The density of lead is 11.3 g/cm^3; the density of gold is 19.3 g/cm^3.

96. Express the following measurements in scientific notation.
  a. 0.000039 kg
  b. 784 L
  c. 0.0830 g
  d. 9,700,000 ng

97. Which of these quantities or relationships are exact?
  a. 10 cm = 1 dm
  b. There are 9 baseball players on the field.
  c. A diamond has a mass of 12.4 g.
  d. The temperature is 21°C.

98. A one-kilogram steel bar is brought to the moon. How are its mass and its weight each affected by this change in location? Explain.

*99. When a piece of copper with a mass of 36.4 g is placed into a graduated cylinder containing 20.00 mL of water, the water level rises to 24.08 mL, completely covering the copper. What is the density of copper?

100. The density of gold is 19.3 g/cm^3. What is the mass, in grams, of a cube of gold that is 2.00 cm on each edge? In kilograms?

*101. A balloon filled with helium will rise upward when released. What does this result show about the relative densities of helium and air?

*102. Explain the difference between the accuracy of a measurement and the precision of a measurement.

103. Give the number of protons and electrons in each of the following:
  a. Cs
  b. Ag
  c. Cd
  d. Se

104. Which of these was an essential part of Dalton's atomic model?
  a. indivisible atoms
  b. electrons
  c. atomic nuclei
  d. neutrons

*105. How do neon-20 and neon-21 differ from each other?

106. The mass of an atom should be very nearly the sum of the masses of its protons and neutrons. The mass of a proton and the mass of a neutron are each very close to 1 amu. Why is the atomic mass of chlorine, 35.453 amu, so far from a whole number?

If You Have Trouble With . . .

Question	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104	105	106
See Chapter	2	2	2	2	3	3	3	3	3	3	3	3	3	3	4	4	4	4

Standardized Test Prep

Select the choice that best answers each question or completes each statement.

Tips for Success

Eliminate Wrong Answers If you don't know which response is correct, start by eliminating those you know are wrong. If you can rule out some choices, you'll have fewer left to consider and you'll increase your chances of choosing the correct answer.

1. Select the correct electron configuration for silicon, atomic number 14.
(A) $1s^22s^22p^23s^23p^23d^24s^2$
(B) $1s^22s^22p^43s^23p^4$
(C) $1s^22s^62p^6$
(D) $1s^22s^22p^63s^23p^2$

2. Which two orbitals have the same shape?
(A) 2*s* and 2*p*
(B) 2*s* and 3*s*
(C) 3*p* and 3*d*
(D) More than one is correct.

3. Which of these statements characterize the nucleus of every atom?
I. It has a positive charge.
II. It is very dense.
III. It is composed of protons, electrons, and neutrons.
(A) I and II only
(B) II and III only
(C) I and III only
(D) I, II, and III

4. As the wavelength of light increases,
(A) the frequency increases.
(B) the speed of light increases.
(C) the energy decreases.
(D) the intensity increases.

5. In the third energy level of an atom,
(A) there are two energy sublevels.
(B) the *f* sublevel has 7 orbitals.
(C) there are three *s* orbitals.
(D) a maximum of 18 electrons are allowed.

The lettered choices below refer to Questions 6–10. A lettered choice may be used once, more than once, or not at all.

(A) s^2p^6 (b) s^2p^2 (C) s^2 (D) s^4p^1 (E) s^2p^4

Which configuration is the configuration of the highest occupied energy level for each of these elements?

6. sulfur

7. germanium

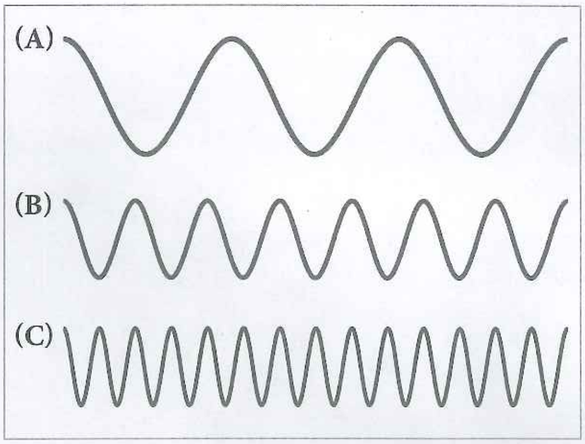
8. beryllium

9. krypton

10. strontium

Use the drawings to answer Questions 11–14. Each drawing represents an electromagnetic wave.

Waves



11. Which wave has the longest wavelength?

12. Which wave has the highest energy?

13. Which wave has the lowest frequency?

14. Which wave has the highest amplitude?

Write a short essay to answer Question 15.

15. Explain the rules that determine how electrons are arranged around the nuclei of atoms.

If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
See Lesson	5.2	5.1	5.1	5.3	5.1	5.2	5.2	5.2	5.2	5.2	5.3	5.3	5.3	5.3	5.2