SECTION 1

The Development of Atomic Theory

Key Ideas	Key Terms	Why It Matters
❯ Who came up with the first theory of atoms? ❯ What did Dalton add to the atomic theory? ❯ How did Thomson discover the electron? ❯ What is Rutherford's atomic model?	**electron** **nucleus**	The electrons stripped from atoms are used in television tubes to help create the images on a television screen.

Atoms are everywhere. They make up the air you are breathing, the chair you are sitting in, and the clothes you are wearing. Atoms determine the properties of matter. For example, the aluminum containers shown in **Figure 1** are lightweight because of the properties of the atoms that make up the aluminum.

The Beginnings of Atomic Theory

Today, it is well known that matter is made up of particles called atoms. But atomic theory was developed slowly over a long period of time. The first theory of atoms was proposed more than 2,000 years ago. ❯ **In the fourth century BCE, the Greek philosopher Democritus suggested that the universe was made of indivisible units.** He called these units *atoms.* "Atom" comes from *atomos,* a Greek word that means "unable to be cut or divided." Democritus thought that movements of atoms caused the changes in matter that he observed.

Democritus did not have evidence for his atomic theory.

Although his theory of atoms explained some observations, Democritus did not have the evidence needed to convince people that atoms existed. Throughout the centuries that followed, some people supported Democritus's theory. But other theories were also proposed.

As the science of chemistry was developing in the 1700s, the emphasis on making careful and repeated measurements in experiments increased. Because of this change, more-precise data were collected and were used to favor one theory over another.

Figure 1 The properties of aluminum containers come from the properties of aluminum atoms, shown magnified here in an image from a scanning tunneling electron microscope.

Dalton's Atomic Theory

In 1808, an English schoolteacher named John Dalton proposed a revised atomic theory. Dalton's theory was developed on a scientific basis, and some parts of his theory still hold true today. Like Democritus, Dalton proposed that atoms could not be divided. › **According to Dalton, all atoms of a given element were exactly alike, and atoms of different elements could join to form compounds.**

Reading Check **How are Dalton's and Democritus's atomic theories similar?** (See Appendix E for answers to Reading Checks.)

Dalton used experimental evidence.

Unlike Democritus, Dalton based his theory on experimental evidence. For instance, scientists were beginning to observe that some substances combined together in consistent ways. According to the *law of definite proportions,* a chemical compound always contains the same elements in exactly the same proportions by weight or mass. For example, any sample of water contains the same proportions of hydrogen and oxygen by mass, as **Figure 2** shows. This and other evidence supported Dalton's theory.

Dalton's theory did not fit all observations.

Today, Dalton's theory is considered the foundation for modern atomic theory. Some parts of Dalton's work turned out to be correct. However, as experiments continued, Dalton's theory could not explain all of the experimental evidence. Like many scientific theories, the atomic theory changed gradually over many years as scientists continued to do experiments and acquire more information.

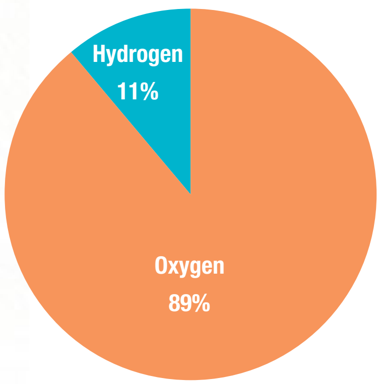
QuickLab 10 min

Evidence for Atoms

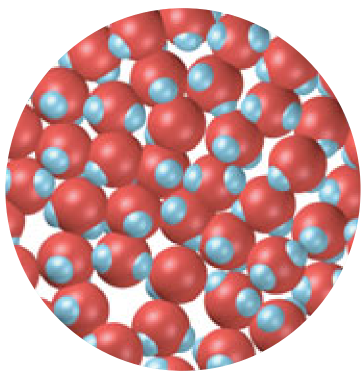
1. Your teacher will provide **two cups** that contain a mixture of **pennies** and **marbles.**
2. Use a **balance** to find the total mass of the pennies in the first cup. Then, find the mass of the marbles.
3. Repeat step 2 for the second cup.
4. Compare the composition of the "compounds" in the two cups in terms of the proportions of marbles and pennies by mass. Do the cups contain the same "compound"?
5. What can you deduce from your results?

Figure 2 **The Law of Definite Proportions**

Water Composition by Mass



For any given water sample, the proportions of hydrogen and oxygen by mass are constant.



This fact suggests that water molecules are made up of atoms that combine in simple whole-number ratios to form compounds.

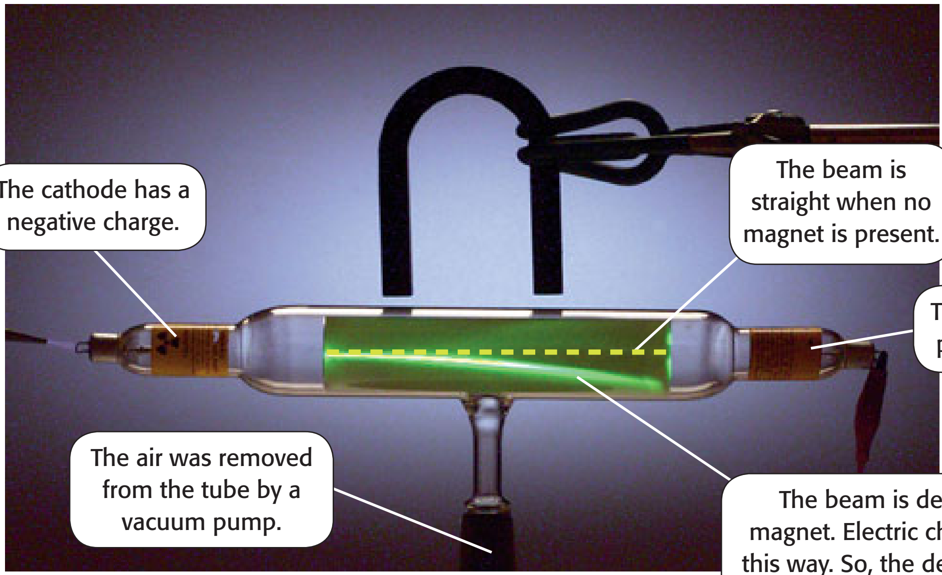


Figure 3 **Thomson's Cathode-Ray Tube Experiment**

Thomson's Model of the Atom

In 1897, J. J. Thomson, a British scientist, conducted an experiment that suggested that atoms were not indivisible. Thomson wasn't planning to learn about the atom. Instead, he was experimenting with electricity. He was studying *cathode rays,* mysterious rays in vacuum tubes. ❯ **Thomson's cathode-ray tube experiment suggested that cathode rays were made of negatively charged particles that came from inside atoms.** This result revealed that atoms could be divided into smaller parts.

Thomson developed the plum-pudding model.

An experiment similar to Thomson's is shown in **Figure 3.** The two metal plates at the ends of the vacuum tube are called the *cathode* and the *anode.* The cathode has a negative charge, and the anode has a positive charge. When a voltage is applied across the plates, a glowing beam comes from the cathode and strikes the anode.

Thomson knew that magnets deflected charges. He reasoned that because all of the air was removed from the tube, the beam must have come from the cathode or from the anode. The direction of the deflection confirmed that the beam was made of negative charges and thus came from the cathode. Thomson had discovered **electrons,** negatively charged particles inside the atom.

Thomson proposed a new model of the atom based on his discovery. In this model, electrons are spread throughout the atom, just as blueberries are spread throughout the muffin in **Figure 4.** Thomson's model, often called the *plum-pudding model,* was named after a dessert that was popular in his day.

electron (ee LEK TRAHN) a subatomic particle that has a negative charge

Figure 4 This blueberry muffin is similar to Thomson's atomic model. **What do the blueberries represent?**

Why It **Matters**

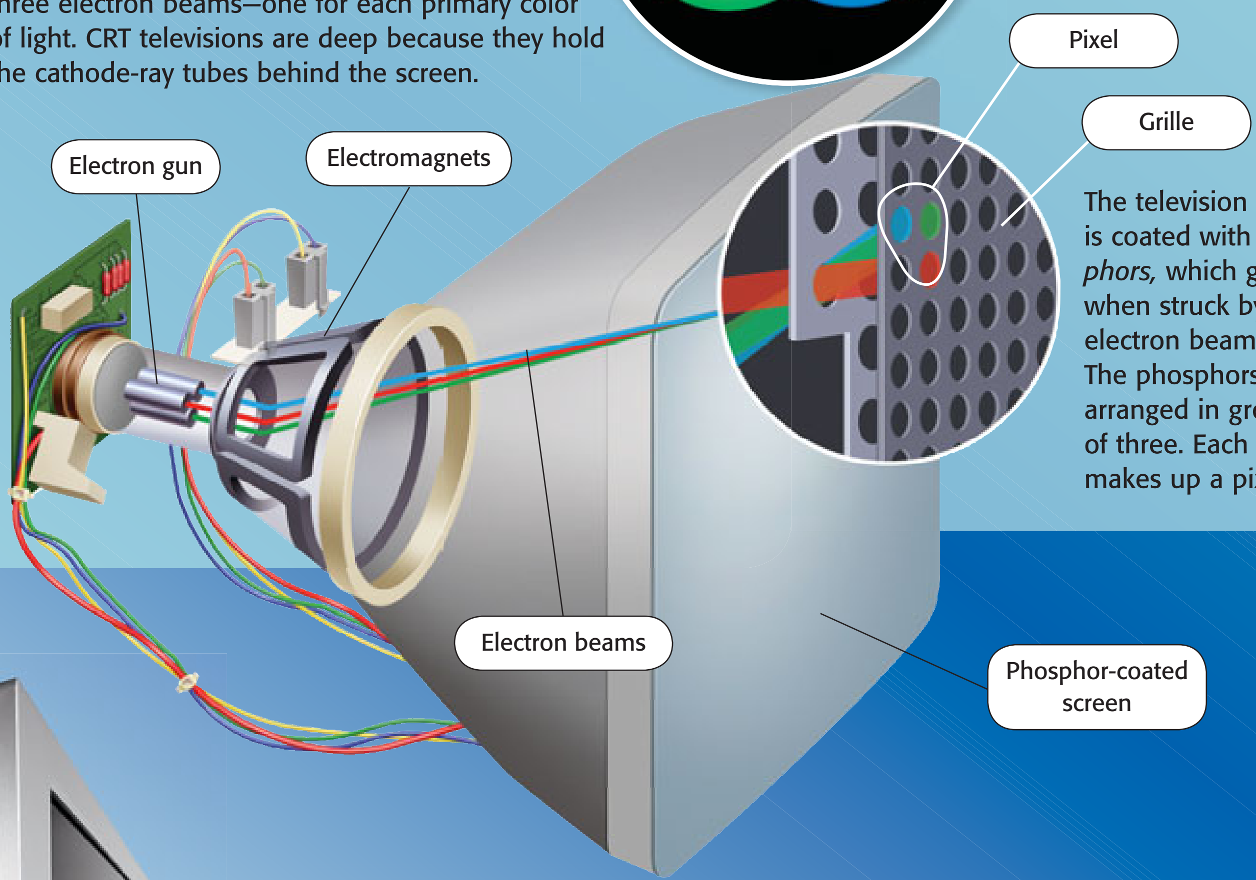
How Do Televisions Work?

REAL WORLD

How does your television screen display images? Television images are made up of thousands of tiny pixels. Your brain puts the pixels together and interprets them as images. How the pixels are created depends on the type of television.

In any television set, red, green, and blue colors combine in different proportions to produce the entire color spectrum.

Some televisions use cathode-ray tubes, like the one in Thomson's experiment. Color TVs contain three electron beams—one for each primary color of light. CRT televisions are deep because they hold the cathode-ray tubes behind the screen.



The television screen is coated with *phosphors,* which glow when struck by the electron beams. The phosphors are arranged in groups of three. Each group makes up a pixel.

Flat-panel televisions use various technologies, including plasmas, to produce pixels. Because these TVs do not contain large cathode-ray tubes, they can be very thin.

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Topic: Television Technology
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YOUR TURN

UNDERSTANDING CONCEPTS

1. How are CRT televisions and flat-panel televisions similar?

ONLINE RESEARCH

2. Use the Internet to learn more about how flat-panel televisions work. Choose one type of flat-panel technology, and create a poster that illustrates how it works.

Rutherford's Model of the Atom

Shortly after Thomson proposed his new atomic model, Ernest Rutherford, another British scientist, developed an experiment to test Thomson's model. Rutherford found that Thomson's model needed to be revised. ❯ **Rutherford proposed that most of the mass of the atom was concentrated at the atom's center.** To understand why Rutherford came to this conclusion, you need to learn about his experiment.

Rutherford conducted the gold-foil experiment.

In Rutherford's experiment, shown in **Figure 5,** two of Rutherford's students aimed a beam of positively charged alpha particles at a very thin sheet of gold foil. In Thomson's model of the atom, the mass and positive charge of an atom are evenly distributed, and electrons are scattered throughout the atom. The positive charge at any location would be too small to affect the paths of the incoming particles. Rutherford predicted that most particles would travel in a straight path and that a few would be slightly deflected.

The observations from the experiment did not match Rutherford's predictions. As **Figure 5** shows, most particles did pass straight through the gold foil, but some were deflected by a large amount. A few particles came straight back. Rutherford wrote, "It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you."

Reading Check **Why were Rutherford's results surprising?**

READING TOOLBOX

Pyramid FoldNote

Create a Pyramid FoldNote to compare the atomic models of Dalton, Thomson, and Rutherford. Be sure to note the similarities and differences between the models.

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Figure 5 **Rutherford's Gold-Foil Experiment**

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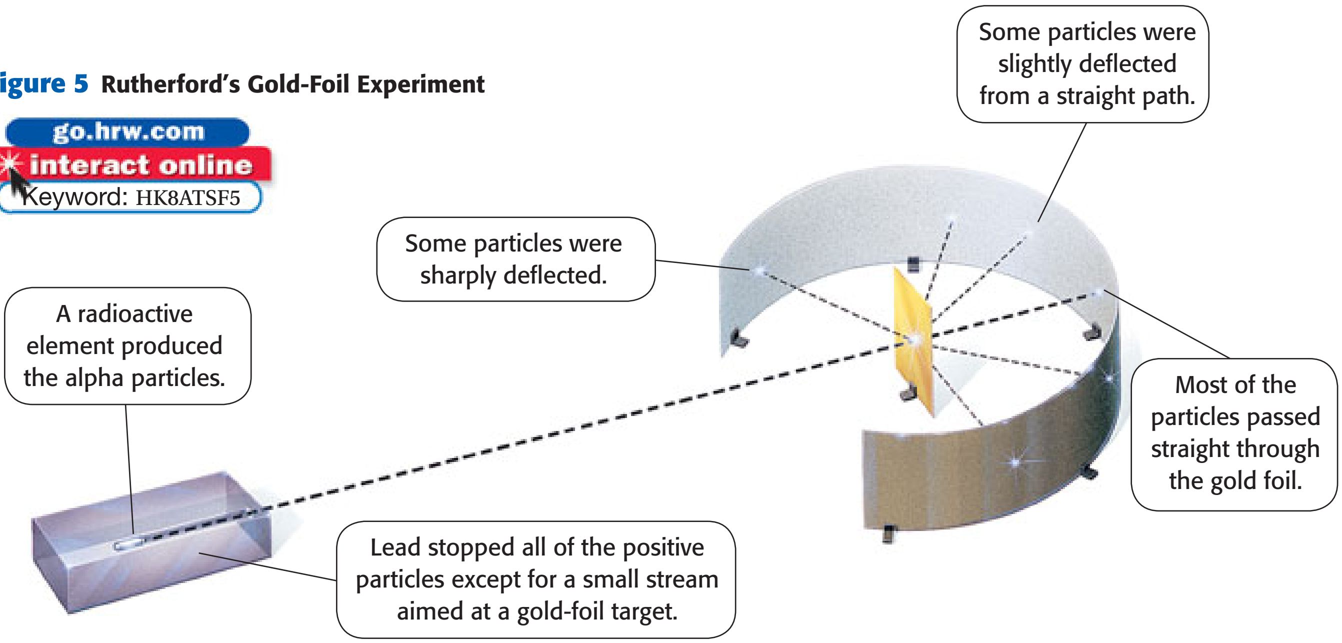


Figure 6 If the nucleus of an atom were the size of a marble, the whole atom would be the size of a football stadium!

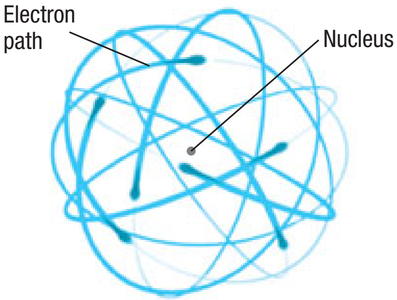
nucleus (NOO klee uhs) an atom's central region, which is made up of protons and neutrons

Rutherford discovered the nucleus.

Rutherford's experiment suggested that an atom's positive charge is concentrated in the center of the atom. This positively charged, dense core of the atom is called the **nucleus.** In the gold-foil experiment, incoming positive charges that passed close to the nucleus were deflected sharply. Incoming positive charges that were aimed directly at the nucleus bounced straight back. Data from Rutherford's experiment suggested that compared with the atom, the nucleus was very small, as **Figure 6** shows.

Rutherford's results led to a new model of the atom, shown in **Figure 7.** In Rutherford's model, negative electrons orbit the positively charged nucleus in much the same way that planets orbit the sun. Today, we understand that the nucleus contains particles called *protons* and *neutrons.* Protons have a positive charge, and neutrons have no charge. You will learn more about these particles in the next section.

Figure 7 In Rutherford's model, electrons orbit the nucleus. (This figure does not accurately represent sizes and distances.) **Is the nucleus positive, negative, or neutral?**



The Rutherford model of the atom

Section 1 Review

KEY IDEAS

1. **Describe** Democritus's atomic theory.
2. **Summarize** the main ideas of Dalton's theory.
3. **Explain** why Dalton's theory was more successful than Democritus's theory.
4. **Compare** Thomson's atomic model with Rutherford's atomic model.

CRITICAL THINKING

5. **Analyzing Experiments**
   a. How did the cathode-ray tube experiment lead to the conclusion that atoms contain electrons?
   b. How did the gold-foil experiment lead to the conclusion that the atom has a nucleus?
6. **Making Inferences** Does the term *indivisible* still describe the atom? Explain.

SECTION 2

The Structure of Atoms

Key Ideas

- What is the difference between protons, neutrons, and electrons?
- What do atoms of an element have in common with other atoms of the same element?
- Why do isotopes of the same element have different atomic masses?
- What unit is used to express atomic mass?

Key Terms

proton
neutron
atomic number
mass number
isotope
unified atomic mass unit
mole

Why It Matters

Radioisotopes emit energy when they decay. To diagnose and treat diseases, doctors use this property of radioisotopes to track where in the body certain atoms go.

Less than 100 years after Dalton published his atomic theory, scientists determined that atoms consisted of smaller particles, such as the electron. In this section, you will learn more about the particles inside the atom.

What Is in an Atom?

Atoms are made up of various subatomic particles. To understand the chemistry of most substances, however, we need to study only three of these particles. **The three main subatomic particles are distinguished by mass, charge, and location in the atom.** **Figure 1** compares these particles.

At the center of each atom is a small, dense *nucleus*. The nucleus is made of **protons,** which have a positive charge, and **neutrons,** which have no charge. Protons and neutrons are almost identical in size and mass. Moving around outside the nucleus is a cloud of very tiny, negatively charged *electrons*. The mass of an electron is much smaller than that of a proton or neutron.

proton (PROH TAHN) a subatomic particle that has a positive charge and that is located in the nucleus of an atom

neutron (NOO TRAHN) a subatomic particle that has no charge and that is located in the nucleus of an atom

Figure 1 **Subatomic Particles**

Particle	Charge	Mass (kg)	Location in the atom
Proton	+1	1.67×10^{-27}	in the nucleus
Neutron	0	1.67×10^{-27}	in the nucleus
Electron	–1	9.11×10^{-31}	outside the nucleus

Figure 2 Helium atoms, including the ones in this helium blimp, are made up of two protons, two neutrons, and two electrons. **Which of these particles defines the element as helium?**

Each element has a unique number of protons.

A hydrogen atom has one proton. A helium atom, shown in **Figure 2,** has two protons. Lithium has three protons. As you move through the periodic table of the elements, this pattern continues. Each element has a unique number of protons. In fact, an element is defined by the number of protons in an atom of that element.

Unreacted atoms have no overall charge.

Even though the protons and electrons in atoms have electric charges, most atoms do not have an overall charge. The reason is that most atoms have an equal number of protons and electrons, whose charges exactly cancel. For example, a helium atom has two protons and two electrons. The atom is neutral because the positive charge of the two protons exactly cancels the negative charge of the two electrons, as shown below.

Charge of two protons:	$+2$
Charge of two neutrons:	0
Charge of two electrons:	-2
Total charge of a helium atom:	0

If an atom gains or loses electrons, it becomes charged. A charged atom is called an *ion.*

Academic Vocabulary

overall (OH vuhr AWL) total; net

The electric force holds the atom together.

Positive and negative charges attract each other with a force known as the *electric force.* Because protons are positive and electrons are negative, protons and electrons are attracted to one another by the electric force. In fact, the electric force between protons in the nucleus and electrons outside the nucleus holds the atom together. On a larger scale, this same force holds solid and liquid materials together. For instance, electric attractions hold water molecules together.

Atomic Number and Mass Number

Atoms of different elements have their own unique structures. Because these atoms have different structures, they have different properties. Atoms of the same element can vary in structure, too. ❯ **Atoms of each element have the same number of protons, but they can have different numbers of neutrons.**

The atomic number equals the number of protons.

The **atomic number** of an element, Z, tells you how many protons are in an atom of the element. Remember that most atoms are neutral because they have an equal number of protons and electrons. Thus, the atomic number also equals the number of electrons in the atom. Because each element is defined by its unique number of protons, each element has a unique atomic number. Hydrogen has only one proton, so $Z = 1$ for hydrogen. The largest naturally occurring element, uranium, has 92 protons, so $Z = 92$ for uranium. The atomic number of a given element never changes.

atomic number (uh TAHM ik NUHM buhr) the number of protons in the nucleus of an atom

mass number (MAS NUHM buhr) the sum of the numbers of protons and neutrons in the nucleus of an atom

The mass number equals the total number of subatomic particles in the nucleus.

The **mass number** of an element, A, equals the number of protons plus the number of neutrons in an atom of the element. A fluorine atom has 9 protons and 10 neutrons, so $A = 19$ for fluorine. Oxygen has 8 protons and 8 neutrons, so $A = 16$ for oxygen. The mass number reflects the number of protons and neutrons (and not the number of electrons) because protons and neutrons provide most of the atom's mass. Although atoms of an element have the same atomic number, they can have different mass numbers because the number of neutrons can vary. **Figure 3** shows which subatomic particles in the nucleus of an atom contribute to the atomic number and which contribute to the mass number.

Reading Check **Which defines an element: the atomic number of the element or the mass number of the element?**

READING TOOLBOX

Pyramid FoldNote

Create a Pyramid FoldNote for the terms *proton, neutron,* and *electron,* and describe which can vary for a given element. Also include the terms *atomic number* and *mass number* in your notes.

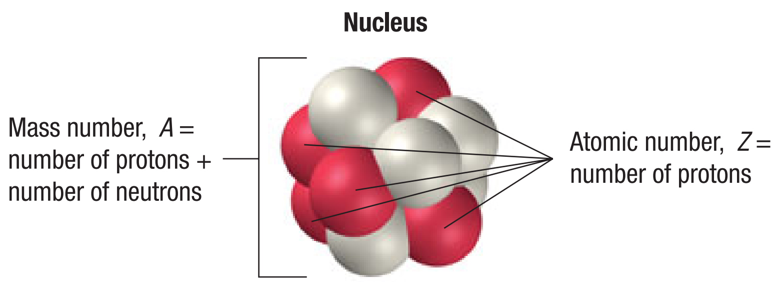


Figure 3 Atoms of the same element have the same number of protons and thus the same atomic number. But because the number of neutrons may vary, atoms of the same element may have different mass numbers.

QuickLab Modeling Isotopes

20 min

Procedure

1. Use **gumdrops of three different colors** to represent protons, neutrons, and electrons. Use **toothpicks** to hold the gumdrops together.
2. Create atomic models for the three isotopes of hydrogen: protium ($A = 1$), deuterium ($A = 2$), and tritium ($A = 3$).
3. Create atomic models for helium-3 ($A = 3$) and helium-4 ($A = 4$).

Analysis

1. How do the isotopes of hydrogen compare to one another?
2. How do the hydrogen isotopes differ from the helium isotopes?
3. Which isotopes have the same mass number? Which isotopes have the same atomic number?

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isotope (IE suh TOHP) an atom that has the same number of protons (or the same atomic number) as other atoms of the same element do but that has a different number of neutrons (and thus a different atomic mass)

Isotopes

As you have learned, atoms of a single element can have different numbers of neutrons and thus different mass numbers. An **isotope** is an atom that has the same number of protons but a different number of neutrons relative to other atoms of the same element. Because they have the same number of protons and electrons, isotopes of an element have the same chemical properties. However, isotopes have different masses. **Isotopes of an element vary in mass because their numbers of neutrons differ.**

Each of the three isotopes of hydrogen, shown in **Figure 4,** has one proton and one electron. The most common hydrogen isotope, protium, does not have any neutrons. Because it has one proton in its nucleus, its mass number, A, is 1. Deuterium, a second isotope of hydrogen, has one neutron as well as one proton in its nucleus. Its mass number, A, is 2. A third isotope, tritium, has two neutrons. Because its nucleus contains two neutrons and one proton, tritium has a mass number of 3.

Reading Check **Which hydrogen isotope has the most mass?**

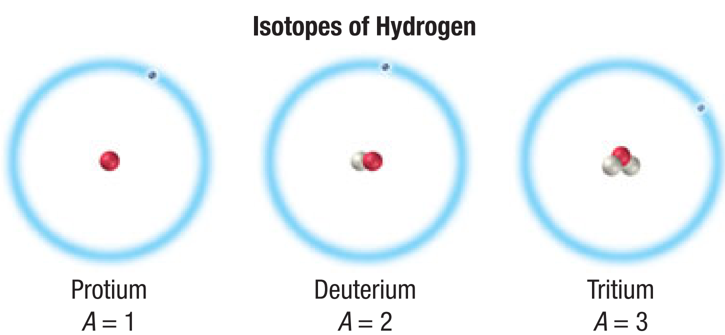

Figure 4 Each isotope of hydrogen has one proton, but the number of neutrons varies. **What is the atomic number, *Z*, of each isotope?**

Some isotopes are more common than others.

Hydrogen is present on Earth and on the sun. In both places, protium is most common. Only a small fraction of the hydrogen found on Earth and on the sun is deuterium. For instance, only 1 out of every 6,000 hydrogen atoms in Earth's crust is a deuterium isotope. Similarly, on the sun, protium isotopes outnumber deuterium isotopes 50,000 to 1.

Tritium is an unstable isotope that decays over time. Thus, tritium is the least common isotope of hydrogen. Unstable isotopes, called *radioisotopes,* emit radiation and decay into other isotopes. A radioisotope continues to decay until the isotope reaches a stable form. Each radioisotope decays at a fixed rate, which can vary from a fraction of a second to millions of years.

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Topic: Radioisotopes
Code: **HK81260**

Why It **Matters**

Nuclear Medicine

SCIENCE & SOCIETY

Radioactive isotopes, or radioisotopes, are widely used in medicine. They are used to diagnose and treat certain conditions. Some isotopes are used to create images similar to X-ray images. Doctors interpret the images to study organ structures and functions. Radioisotopes are also used to study organ metabolisms and to identify and treat cancer.

In the full-body image shown here, the radioisotope *technetium-99m*, along with a biological agent that localizes the radioactivity in the bones, was injected into the body. The image, called a *colored gamma scan* or *scintigram,* shows a healthy human skeleton.

A radioisotope that has been injected into the body emits small amounts of radiation. A special camera, such as the one shown here, detects the radiation. A computer uses the information from the camera to create the image. The image is often interpreted by a radiologist, a doctor who specializes in imaging technologies.

YOUR TURN

CRITICAL THINKING

1. Will radioisotopes injected into the body remain in the body forever? Explain.

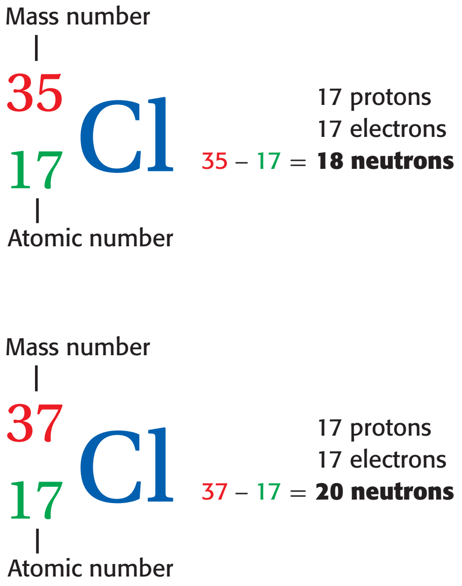


Figure 5 One isotope of chlorine has 18 neutrons, while the other isotope has 20 neutrons.

The number of neutrons can be calculated.

To represent different isotopes, you can write the mass number and atomic number of the isotope before the symbol of the element. The two isotopes of chlorine are represented this way in **Figure 5.** (Sometimes the atomic number is omitted because it is always the same for any given element.)

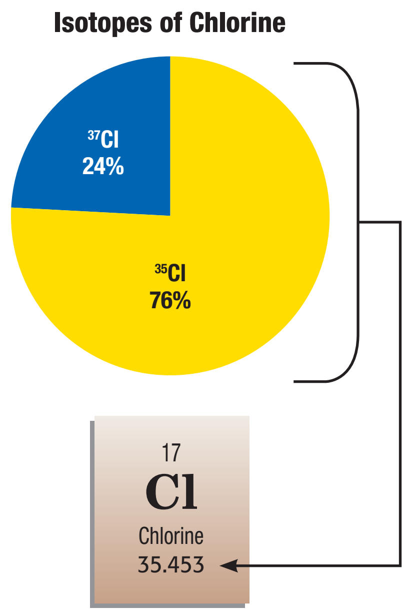
If you know the atomic number and mass number, you can calculate the number of neutrons that an atom has. For example, the isotope of uranium that is used in nuclear reactors is uranium-235, or $^{235}_{92}U$. Like all uranium atoms, this isotope has an atomic number of 92, so it has 92 protons. Its mass number is 235, so there are a total of 235 protons and neutrons. The number of neutrons can be found by subtracting the atomic number from the mass number, as shown below.

Mass number (A):	235
Atomic number (Z):	− 92
Number of neutrons:	143

Atomic Masses

The mass of a single atom is very small. The mass of a single fluorine atom is less than one trillionth of a billionth of a gram. **Because working with such tiny masses is difficult, atomic masses are usually expressed in unified atomic mass units.** A **unified atomic mass unit** (u) is equal to one-twelfth of the mass of a carbon-12 atom. (This unit is sometimes called the atomic mass unit, amu.) Carbon-12, an isotope of carbon, has six protons and six neutrons, so each individual proton and neutron has a mass of about 1.0 u. Recall that electrons contribute very little mass to an atom.

Figure 6 The average atomic mass of chlorine is closer to 35 u than it is to 37 u. **Which chlorine isotope is more common?**



Note: Calculations using the values from the pie graph do not give a result of exactly 35.453 u because of rounding.

Average atomic mass is a weighted average.

Often, the atomic mass listed for an element in the periodic table is an average atomic mass for the element as found in nature. The *average atomic mass* for an element is a weighted average. In other words, commonly found isotopes have a greater effect on the average atomic mass than rarely found isotopes do.

For example, **Figure 6** shows how the natural abundance of chlorine's two isotopes affects chlorine's average atomic mass, which is 35.453 u. This mass is closer to 35 u than it is to 37 u. The reason is that the atoms of chlorine that have a mass of nearly 35 u are more common in nature. Thus, they make a greater contribution to chlorine's average atomic mass.

The mole is useful for counting small particles.

Because chemists often deal with large numbers of small particles, they use a large counting unit called the **mole** (mol). A mole is a collection of a very large number of particles.

1 mol = 602, 213, 670, 000, 000, 000, 000, 000 particles

This number is usually written as 6.022×10^{23} and is called *Avogadro's number.* The number is named for Italian scientist Amedeo Avogadro. Why is 6.022×10^{23} the number of particles in one mole? The mole has been defined as the number of atoms in 12.00 grams of carbon-12. Experiments have shown this value to be 6.022×10^{23}. So, one mole of a substance contains 6.022×10^{23} particles of that substance.

The following example demonstrates the magnitude of Avogadro's number: 6.022×10^{23} popcorn kernels would cover the United States to form a pile about 500 km (310 mi) tall! So, Avogadro's number is not useful for counting items such as popcorn kernels but is useful for counting atoms or molecules.

Reading Check **How many particles are in 1 mol of iron?**

Moles and grams are related.

The mass in grams of one mole of a substance is called *molar mass.* For example, 1 mol of carbon-12 atoms has a mass of 12.00 g, so the molar mass of carbon-12 is 12.00 g/mol. **Figure 7** shows the molar mass of magnesium.

In nature, elements often occur as mixtures of isotopes. So, a mole of an element usually contains several isotopes. As a result, an element's molar mass in grams per mole equals its average atomic mass in unified atomic mass units, u. The average atomic mass of carbon is 12.01 u. So, one mole of carbon has a mass of 12.01 g. Because this mass is a weighted average of the masses of several isotopes of carbon, it differs from the molar mass of carbon-12, which is a single isotope.

Integrating Space Science

Counting Stars How many stars are in the universe? In 2003, a group of astronomers estimated that the visible universe—the portion that our telescopes can reach—contains 70,000,000,000,000,000,000,000 stars, or 7×10^{22} stars. This quantity is about one-tenth of Avogadro's number! Of course, this estimate is based on observations from existing telescopes. As improvements in telescopes occur, the estimate could increase. Perhaps the estimate will reach Avogadro's number someday.

unified atomic mass unit (YOON uh FIED uh TAHM ik MAS YOON it) a unit of mass that describes the mass of an atom or molecule; it is exactly 1/12 the mass of a carbon atom with mass number 12 (symbol, u)

mole (MOHL) the SI base unit used to measure the amount of a substance whose number of particles is the same as the number of atoms of carbon in exactly 12 g of carbon-12

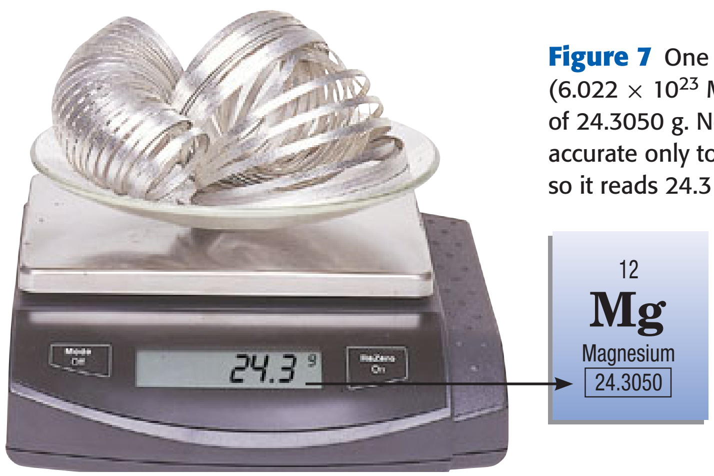


Figure 7 One mole of magnesium (6.022×10^{23} Mg atoms) has a mass of 24.3050 g. Note that the balance is accurate only to one-tenth of a gram, so it reads 24.3 g.

You can convert between moles and grams.

Converting between the amount of an element in moles and the mass of an element in grams is outlined in **Figure 8.** For example, to determine the mass of 5.50 mol of iron, first you must find iron, Fe, in the periodic table. The average atomic mass of iron, rounded to the hundredths place, is 55.84 u. So, the molar mass of iron is 55.84 g/mol. Next, you must set up the problem by using the molar mass as if it were a conversion factor, as shown below.

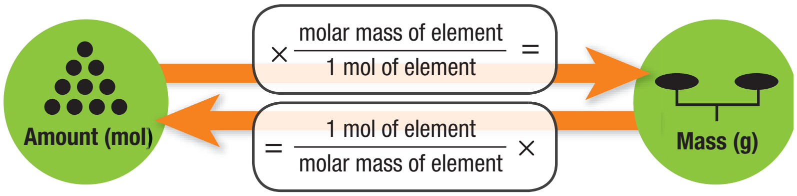


Figure 8 The molar mass of an element allows you to convert between the amount of the element in moles and the mass of the element in grams.

Math Skills: Converting Moles to Grams

Determine the mass in grams of 5.50 mol of iron.

Identify

List the given and unknown values.

Given:

amount of iron = 5.50 mol

molar mass of iron = 55.84 g/mol

Unknown:

mass of iron = ? g

Plan

Write down the conversion factor that converts moles to grams.

The conversion factor you choose should have what you are trying to find (grams of Fe) in the numerator and what you want to cancel (moles of Fe) in the denominator.

$$\frac{55.84 \text{ g Fe}}{1 \text{ mol Fe}}$$

Solve

Multiply the amount of iron by this conversion factor, and solve.

$$5.50 \cancel{\text{ mol Fe}} \times \frac{55.84 \text{ g Fe}}{1 \cancel{\text{ mol Fe}}} = 307 \text{ g Fe}$$

Practice

1. What is the mass in grams of each of the following?
   - **a.** 2.50 mol of sulfur, S
   - **b.** 1.80 mol of calcium, Ca
   - **c.** 0.50 mol of carbon, C
   - **d.** 3.20 mol of copper, Cu

For more practice, visit go.hrw.com and enter keyword **HK8MP.**

Practice Hint

- Remember to use the periodic table to find molar masses. The average atomic mass of an element is equal to the molar mass of the element. For consistency, this book rounds values to the hundredths place.
- Follow the procedure shown in the sample to convert grams to moles, but be sure to reverse the conversion factor, as shown in Figure 8. You can convert grams to moles to check your answers to the practice questions.

Compounds also have molar masses.

Recall that compounds are made up of atoms joined together in specific ratios. To find the molar mass of a compound, you can add up the molar masses of all of the atoms in a molecule of the compound. For example, to find the molar mass of water, H_2O, first you must find the masses of hydrogen and oxygen in the periodic table. Oxygen's average atomic mass is 16.00 u, so its molar mass is 16.00 g/mol. The molar mass of hydrogen is 1.01 g/mol. You must multiply this value by 2 because a water molecule contains two hydrogen atoms. Thus, the molar mass of H_2O can be calculated as follows:

$$\text{molar mass of } H_2O = (2 \times 1.01 \text{ g/mol}) + 16.00 \text{ g/mol} = 18.02 \text{ g/mol}$$

What does this value tell you? As you learned earlier, molar mass equals the mass in grams of 1 mol of a substance. Thus, the mass of 1 mol of water is 18.02 g. In other words, the total mass of 6.022×10^{23} water molecules is 18.02 g. Take a look at **Figure 9.** Then use information in the caption to find the molar mass of carbon monoxide, another common compound.

Figure 9 Burning charcoal produces carbon dioxide and carbon monoxide, CO, which is a colorless, odorless gas. **What is the molar mass of CO?**

Section 2 Review

KEY IDEAS

1. **List** the charge, mass, and location of each of the three subatomic particles found in atoms.
2. **Explain** how you can use an atom's mass number and atomic number to determine the number of protons, electrons, and neutrons in the atom.
3. **Identify** the subatomic particle used to define an element, and explain why this particle is used.
4. **Explain** why the masses of atoms of the same element may differ.
5. **Calculate** the number of neutrons that each of the following isotopes contains. Use the periodic table to find the atomic numbers.
   - **a.** carbon-14
   - **b.** nitrogen-15
   - **c.** sulfur-35
   - **d.** calcium-45
6. **Identify** the unit that is used for atomic masses.
7. **Define** Avogadro's number, and describe how it relates to a mole of a substance.
8. **Determine** the molar mass of each of the following elements:
   - **a.** manganese, Mn
   - **b.** cadmium, Cd
   - **c.** arsenic, As
   - **d.** strontium, Sr

CRITICAL THINKING

9. **Making Inferences** If an atom loses electrons, will it have an overall charge?
10. **Making Predictions** Predict which isotope of nitrogen is more common in nature: nitrogen-14 or nitrogen-15. (Hint: What is the average atomic mass listed for nitrogen in the periodic table?)
11. **Applying Concepts** Which has a greater number of atoms: 3.0 g of iron, Fe, or 2.0 g of sulfur, S?
12. **Drawing Conclusions** A graph of the amount of a particular element in moles versus the mass of the element in grams is a straight line. Explain why.

Math Skills

13. What is the mass in grams of 0.48 mol of platinum, Pt?
14. What is the mass in grams of 3.1 mol of mercury, Hg?
15. How many moles does 11 g of silicon, Si, contain?
16. How many moles does 205 g of helium, He, contain?

SECTION 3

Modern Atomic Theory

Key Ideas

- What is the modern model of the atom?
- How are the energy levels of an atom filled?
- What makes an electron jump to a new energy level?

Key Terms

orbital
valence electron
photon

Why It Matters

Modern atomic theory explains why different elements produce different colors in fireworks.

Dalton's theory that the atom could not be split had to be modified after the discovery that atoms are made of protons, neutrons, and electrons. Like most scientific models and theories, the model of the atom has been revised many times to explain new discoveries.

Modern Models of the Atom

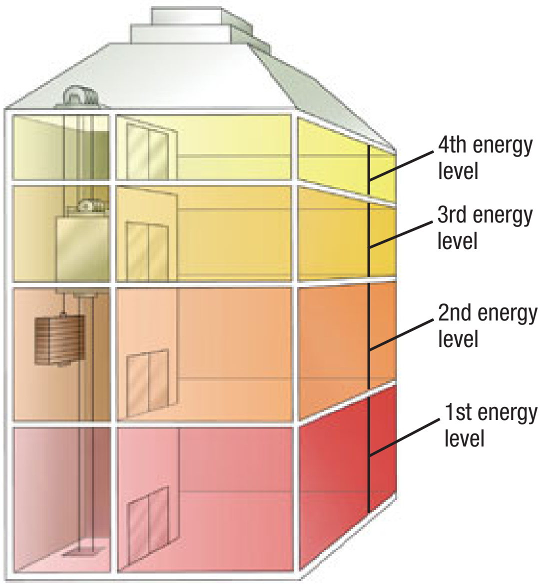
The modern model of the atom, which accounts for many new discoveries in the early 20th century, is very different from the model proposed by Rutherford. **In the modern atomic model, electrons can be found only in certain energy levels, not between levels. Furthermore, the location of electrons cannot be predicted precisely.** This model is not as easy to visualize as the earlier models that you have studied.

Electron location is limited to energy levels.

In 1913, Niels Bohr, a Danish physicist, suggested that the energy of each electron was related to the electron's path around the nucleus. Electrons can be in only certain energy levels. They must gain energy to move to a higher energy level or must lose energy to move to a lower energy level. Bohr's description of energy levels is still used by scientists today.

One way to imagine Bohr's model is to compare an atom to the stairless building shown in **Figure 1.** Imagine that the nucleus is in a deep basement. The energy levels begin on the first floor, above the basement. Electrons can be on any floor, but they cannot be between floors. Electrons gain energy by riding up the elevator and lose energy by riding down.

Figure 1 The energy levels of an atom are like the floors of the building shown here. The energy difference between energy levels decreases as the energy level increases.



Electrons act like waves.

By 1925, Bohr's model of the atom no longer explained all aspects of electron behavior. A new model, which no longer assumed that electrons orbited the nucleus along definite paths in the same way that planets orbit the sun, was proposed. According to this new atomic model, electrons behave more like waves on a vibrating string than like particles.

Reading Check **How does the electron-wave model of the atom differ from earlier atomic models?**

Academic Vocabulary

assume (uh SOOM) to accept without proof

READING TOOLBOX

Making Comparisons

As you read this section, make a list of comparisons that you find. For each pair of items, write down ways in which the two items are alike and ways in which they differ.

The exact location of an electron cannot be determined.

Imagine the moving propeller of a plane, such as the one shown in **Figure 2.** If you were asked to identify the location of any of the blades at a certain instant, you would not be able to give an exact answer. Knowing the exact location of any of the blades is very difficult because the blades are moving so quickly. All you know is that each blade could be anywhere within the blurred area that you see as the blades turn.

Similarly, determining the exact location of an electron in an atom and the speed and direction of the electron is impossible. The best that scientists can do is to calculate the chance of finding an electron in a certain place within an atom. One way to show visually the likelihood of finding an electron in a given location is by shading. The darker the shading, the better the chance of finding an electron at that location. The shaded region is called an **orbital.**

orbital (AWR buh tuhl) a region in an atom where there is a high probability of finding electrons

Figure 2 The exact location of any of the blades of this airplane is difficult to determine. Likewise, pinpointing the location of an electron in an atom is difficult.

The shaded region, or orbital, shows where electrons are most likely to be.

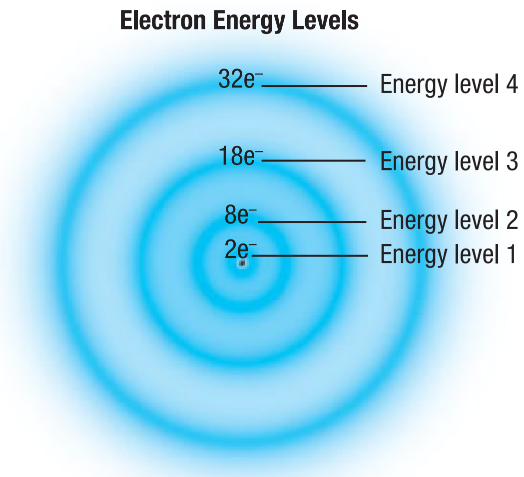


Figure 3 Each energy level holds a certain number of electrons. **How many total electrons can the first four energy levels hold?**

SCILINKS.
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Topic: Atomic Orbitals
Code: **HK80118**

Electron Energy Levels

Within the atom, electrons that have various amounts of energy exist in different energy levels. There are many possible energy levels that an electron can occupy. **Figure 3** shows how the first four energy levels of an atom are filled. **The number of energy levels that are filled in an atom depends on the number of electrons.** For example, a lithium atom has three electrons: two in the first energy level and one in the second.

The electrons in the outer energy level of an atom are called **valence electrons.** Valence electrons determine the chemical properties of an atom. Because lithium has one electron in its outer energy level, it has one valence electron.

There are four types of orbitals.

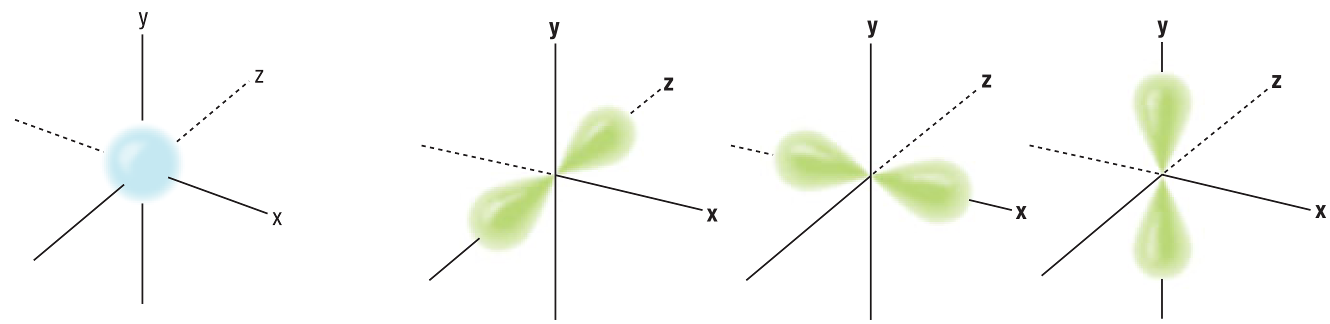
Within each energy level, electrons occupy orbitals that have the lowest energy. There are four kinds of orbitals: s, p, d, and f orbitals. **Figure 4** shows the s and p orbitals.

The s orbital is the simplest kind of orbital. An s orbital has only one possible orientation in space because it is shaped like a sphere, as the figure shows. An s orbital has the lowest energy and can hold two electrons.

A p orbital, on the other hand, is shaped like a dumbbell and can be oriented in space in one of three ways. The axes on the graphs in **Figure 4** can help you picture how these orbitals look in three dimensions. Imagine that the *y*-axis is flat on the page. Imagine that the dotted lines on the *x*- and *z*-axes are going into the page and the darker lines are coming out of the page. Because each p orbital can hold two electrons, the three p orbitals can hold a total of six electrons.

The d and f orbitals are much more complex. There are five possible d orbitals and seven possible f orbitals. Although all of these orbitals are very different in shape, each can hold a maximum of two electrons.

Figure 4 **The s and p Orbitals**



An s orbital is shaped like a sphere, so it has only one possible orientation in space. An s orbital can hold a maximum of two electrons.

Each of these p orbitals can hold a maximum of two electrons, so all three p orbitals can hold a total of six electrons.

Orbitals determine the number of electrons that each level can hold.

You have seen that each energy level contains a certain number of electrons. The orbitals in each energy level determine the total number of electrons that the level can hold, as **Figure 5** shows. For instance, you learned earlier that the second energy level can hold eight electrons. The reason is that this level contains four orbitals: one s orbital and three p orbitals. Because each orbital can hold two electrons, the second energy level holds $4 \times 2 = 8$ electrons.

Figure 5 Orbitals and Electrons for Energy Levels 1–4

Energy level	Number of orbitals by type				Total number of orbitals		Number of electrons
	s	p	d	f			
1	1				$1 = 1$	$\times 2 =$	2
2	1	3			$1 + 3 = 4$		8
3	1	3	5		$1 + 3 + 5 = 9$		18
4	1	3	5	7	$1 + 3 + 5 + 7 = 16$		32

Electron Transitions

As you have learned, the modern model of the atom limits the location of electrons to specific energy levels. An electron is never found between these levels. Instead, it "jumps" from one level to the next. What makes an electron move from one level to another? ❯ **Electrons jump between energy levels when an atom gains or loses energy.**

The lowest state of energy of an electron is called the *ground state.* At normal temperatures, most electrons are in the ground state. However, if an electron gains energy, it moves to an *excited state.* It gains energy by absorbing a particle of light, called a **photon.** The electron may then fall back to a lower level. In doing so, the electron releases a photon.

Photons of light have different energies. The energy of a photon determines to which level the electron will jump. Think back to the elevator analogy. When an electron absorbs a photon, it gains energy and "rides up the elevator." The energy of the photon determines the level up to which the electron rides. When the electron loses energy and "rides down the elevator," a photon is released. In this case, the change in levels determines the energy of the emitted photon. Electrons can move between any two levels of the atom.

Reading Check What makes an electron jump from the ground state to an excited state?

QuickLab 20 min

Electron Levels

1. Make a table that has 10 columns and 4 rows. Label the first cell "Energy level." Label the remaining cells in the first row "s" (1 cell), "p" (3 cells), and "d" (5 cells). Label the remaining cells in the first column "1," "2," and "3."
2. For each energy level, place an *X* in cells that correspond to orbitals that are not found in that level.
3. Place **pennies** in empty cells to represent electrons. Each "orbital" can hold two "electrons."
4. Draw Bohr models of atoms whose atomic numbers are 3, 5, 10, and 20. For each model, fill in the cells with the correct number of pennies (in order) to see how the electrons are located in the energy levels.

valence electron (VAY luhns ee LEK TRAHN) an electron that is found in the outermost shell of an atom and that determines the atom's chemical properties

photon (FOH TAHN) a unit or quantum of light

Integrating Space Science

Spectral Analysis Astronomers use the principle that every element emits a unique set of wavelengths to learn about elements in space. This method, called *spectral analysis,* has been used to identify elements in stars. The element helium was discovered through spectral analysis of light from the sun. It was named *helium* because the Greek word for "sun" is *helios.* Helium was later found on Earth.

Atoms absorb or emit light at certain wavelengths.

You have learned that photons are particles of light. The energy of a photon is related to the wavelength of the light. High-energy photons have short wavelengths, and low-energy photons have long wavelengths. Because atoms gain or lose energy in certain amounts, they can absorb or emit only certain wavelengths.

Because each element has a unique atomic structure, the wavelengths emitted depend on the particular element. For instance, the set of wavelengths emitted by hydrogen differs from the set of wavelengths emitted by any other element. For this reason, the wavelengths can be used to identify the substance. They are a type of "atomic fingerprint."

The emission of photons produces the light in neon signs. The wavelength of visible light determines the color of the light. For instance, the wavelengths emitted in neon gas produce the red color shown in **Figure 6.** Gases of other elements produce other colors.

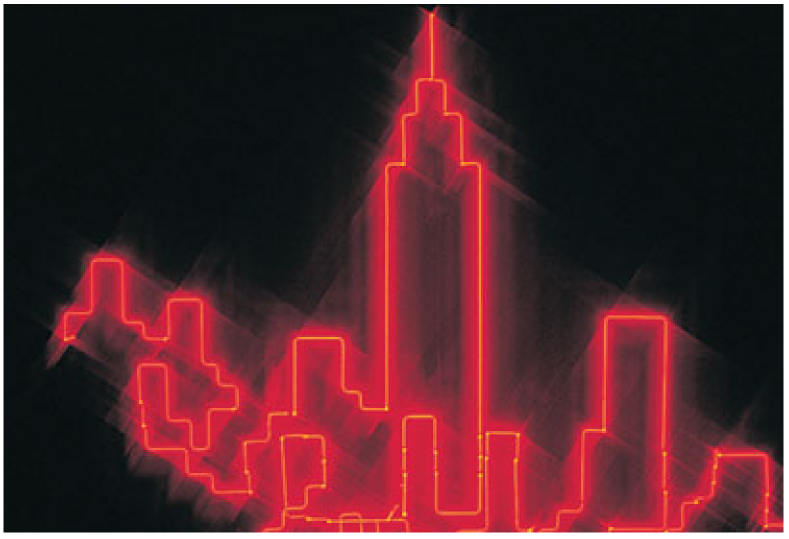


Figure 6 This neon sign lights up because atoms first gain energy from electricity and then release this energy in the form of light. **What happens to the electrons as the light is released?**

Section 3 Review

KEY IDEAS

1. **State** two key features of the modern model of the atom.
2. **Explain** what determines how the energy levels in an atom are filled.
3. **Describe** what happens when an electron jumps from one energy level to another.
4. **Identify** how many electrons the third energy level can hold, and explain why this is the case.

CRITICAL THINKING

5. **Analyzing Models** Compare an atom's structure to a ladder. Identify one way in which a ladder is not a good model for the atom.
6. **Making Comparisons** Explain how Bohr's model and the modern model of the atom differ in terms of the path of an electron.
7. **Applying Concepts** How many valence electrons does nitrogen ($Z = 7$) have?

Why It **Matters**

REAL WORLD

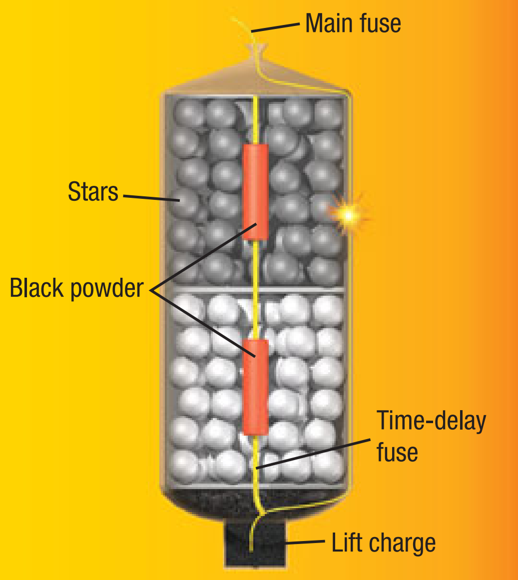
How Do Fireworks Work?

Today's firework displays often feature complex patterns and vivid colors. The colors in fireworks are produced by the emission of photons. The photon wavelength—and thus the color—depends on the compounds that are used. Just as neon gas produces red light, various compounds create specific firework colors. For instance, sodium salts produce yellow, and magnesium and aluminum produce white. These coloring agents are packed into the "stars" inside the shell, often by hand.

4 The second shell continues to shoot into the air. When the time-delay fuse reaches the black powder in this shell, the second set of fireworks explodes.

3 The time-delay fuse inside the shell eventually ignites the black powder in the lower shell, which creates the first set of fireworks.

2 The main fuse quickly ignites the lift charge, which shoots the firework into the air. At the same time, the time-delay fuse begins burning slowly. This system of fuses creates two sets of fireworks from this multibreak shell.



1 At the push of a button on a master control board, the main fuse ignites and moves down toward the lift charge. When ignited, the stars in each shell burst into dazzling displays of color. Black powder, or gunpowder, is used as the explosive.

SCILINKS.
www.scilinks.org
Topic: Fireworks
Code: **HK81697**

YOUR TURN

UNDERSTANDING CONCEPTS

1. How is a firework similar to a neon sign?

ONLINE RESEARCH

2. Use the Internet to learn about the history of fireworks. Create a visual timeline based on the results of your research.

CHAPTER 4

Summary

go.hrw.com
SUPER SUMMARY
KEYWORD: HK8ATSS

Key Ideas

Section 1 The Development of Atomic Theory

- **The Beginnings of Atomic Theory** Democritus suggested that the universe was made of indivisible units called *atoms*. (p. 113)
- **Dalton's Atomic Theory** According to Dalton, all atoms of a given element were exactly alike, and atoms of different elements could join to form compounds. (p. 114)
- **Thomson's Model of the Atom** Thomson's cathode-ray tube experiment suggested that cathode rays were made of negatively-charged particles that came from inside atoms. (p. 115)
- **Rutherford's Model of the Atom** Rutherford proposed that most of the mass of the atom was concentrated at the atom's center. (p. 117)

Key Terms

electron, p. 115
nucleus, p. 118

Section 2 The Structure of Atoms

- **What Is in an Atom?** The three main subatomic particles are distinguished by mass, charge, and location in the atom. (p. 119)
- **Atomic Number and Mass Number** Atoms of an element have the same number of protons, but they can have different numbers of neutrons. (p. 121)
- **Isotopes** Isotopes of an element vary in mass because their numbers of neutrons differ. (p. 122)
- **Atomic Masses** Atomic masses are usually expressed in unified atomic mass units. (p. 124)

Key Terms

proton, p. 119
neutron, p. 119
atomic number, p. 121
mass number, p. 121
isotope, p. 122
unified atomic mass unit, p. 124
mole, p. 125

Section 3 Modern Atomic Theory

- **Modern Models of the Atom** Electrons can be found only in certain energy levels. The location of electrons cannot be predicted precisely. (p. 128)
- **Electron Energy Levels** The number of energy levels that are filled in an atom depends on the number of electrons. (p. 130)
- **Electron Transitions** Electrons jump between levels when an atom gains or loses energy. (p. 131)

Key Terms

orbital, p. 129
valence electron, p. 130
photon, p. 131