Section 4.1

Objectives

- **Compare and contrast** the atomic models of Democritus, Aristotle, and Dalton.
- **Understand** how Dalton's theory explains the conservation of mass.

Review Vocabulary

theory: an explanation supported by many experiments; is still subject to new experimental data, can be modified, and is considered successful if it can be used to make predictions that are true

New Vocabulary

Dalton's atomic theory

Early Ideas About Matter

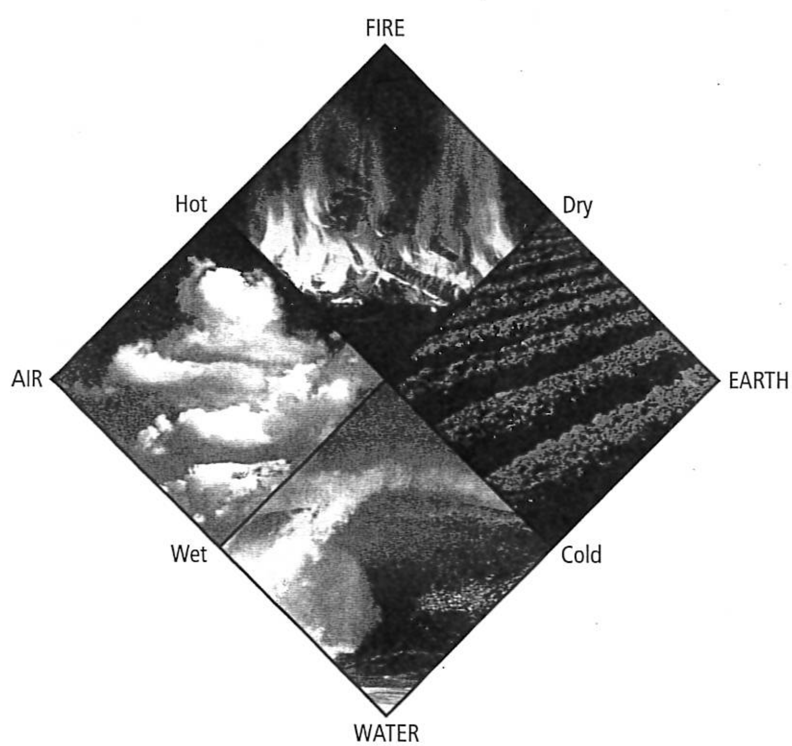
MAIN Idea The ancient Greeks tried to explain matter, but the scientific study of the atom began with John Dalton in the early 1800s.

Real-World Reading Link A football team might practice and experiment with different plays in order to develop the best-possible game plan. As they see the results of their plans, coaches can make adjustments to refine the team's play. Similarly, scientists over the last 200 years have experimented with different models of the atom, refining their models as they collected new data.

Greek Philosophers

Science as we know it today did not exist several thousand years ago. No one knew what a controlled experiment was, and there were few tools for scientific exploration. In this setting, the power of the mind and intellectual thought were considered the primary avenues to the truth. Curiosity sparked the interest of scholarly thinkers known as philosophers who considered the many mysteries of life. As they speculated about the nature of matter, many of the philosophers formulated explanations based on their own life experiences.

Many of them concluded that matter was composed of things such as earth, water, air, and fire, as shown in **Figure 4.1.** It was also commonly accepted that matter could be endlessly divided into smaller and smaller pieces. While these early ideas were creative, there was no method available to test their validity.



■ **Figure 4.1** Many Greek philosophers thought that matter was composed of four elements: earth, air, water, and fire. They also associated properties with each element. The pairing of opposite properties, such as hot and cold, and wet and dry, mirrored the symmetry they observed in nature. These early ideas were incorrect and non-scientific.

Democritus The Greek philosopher Democritus (460–370 B.C.) was the first person to propose the idea that matter was not infinitely divisible. He believed matter was made up of tiny individual particles called *atomos*, from which the English word *atom* is derived. Democritus believed that atoms could not be created, destroyed, or further divided. Democritus and a summary of his ideas are shown in **Table 4.1.**

While a number of Democritus's ideas do not agree with modern atomic theory, his belief in the existence of atoms was amazingly ahead of his time. However, his ideas were met with criticism from other philosophers who asked, "What holds the atoms together?" Democritus could not answer the question.

VOCABULARY
WORD ORIGIN
Atom
comes from the Greek word *atomos*, meaning *indivisible*

Aristotle Other criticisms came from Aristotle (384–322 B.C.), one of the most influential Greek philosophers. He rejected the notion of atoms because it did not agree with his own ideas about nature. One of Aristotle's major criticisms concerned the idea that atoms moved through empty space. He did not believe that empty space could exist. His ideas are also presented in **Table 4.1.** Because Aristotle was one of the most influential philosophers of his time, Democritus's atomic theory was eventually rejected.

In fairness to Democritus, it was impossible for him or anyone else of his time to determine what held the atoms together. More than two thousand years would pass before scientists would know the answer. However, it is important to realize that Democritus's ideas were just that—ideas, not science. Without the ability to conduct controlled experiments, Democritus could not test the validity of his ideas.

Unfortunately for the advancement of science, Aristotle was able to gain wide acceptance for his ideas on nature—ideas that denied the existence of atoms. Incredibly, the influence of Aristotle was so great and the development of science so primitive that his denial of the existence of atoms went largely unchallenged for two thousand years!

Reading Check **Infer** why it was hard for Democritus to defend his ideas.

Table 4.1 Ancient Greek Ideas About Matter

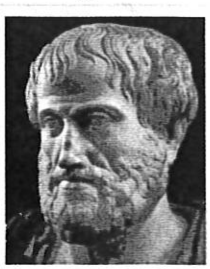
Philosopher		Ideas
Democritus (460–370 B.C.)		• Matter is composed of atoms, which move through empty space. • Atoms are solid, homogeneous, indestructible, and indivisible. • Different kinds of atoms have different sizes and shapes. • Size, shape, and movement of atoms determine the properties of matter.
Aristotle (384–322 B.C.)		• Empty space cannot exist. • Matter is made of earth, fire, air, and water.

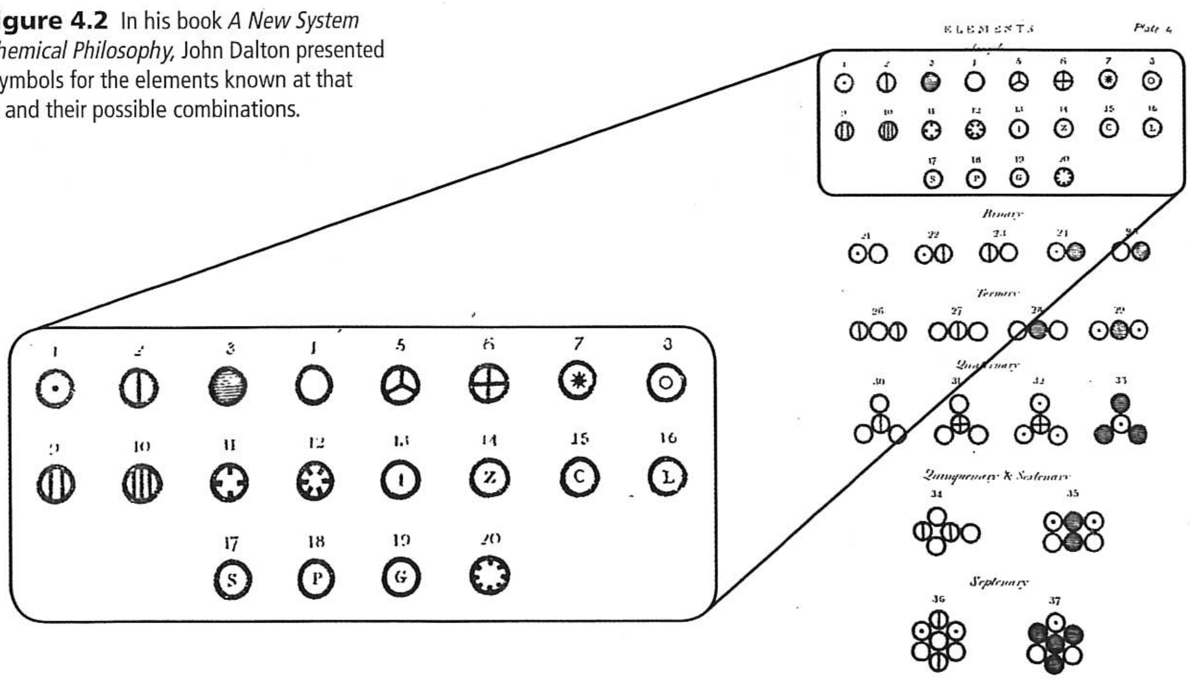
Table 4.2 Dalton's Atomic Theory

Scientist	Ideas
Dalton **(1766–1844)** 	• Matter is composed of extremely small particles called atoms. • Atoms are indivisible and indestructible. • Atoms of a given element are identical in size, mass, and chemical properties. • Atoms of a specific element are different from those of another element. • Different atoms combine in simple whole-number ratios to form compounds. • In a chemical reaction, atoms are separated, combined or rearranged.

John Dalton Although the concept of the atom was revived in the eighteenth century, it took another hundred years before significant progress was made. The work done in the nineteenth century by John Dalton (1766–1844), a schoolteacher in England, marks the beginning of the development of modern atomic theory. Dalton revived and revised Democritus's ideas based on the results of scientific research he conducted. In many ways, Democritus's and Dalton's ideas are similar.

Thanks to advancements in science since Democritus's day, Dalton was able to perform experiments that allowed him to refine and support his hypotheses. He studied numerous chemical reactions, making careful observations and measurements along the way. He was able to determine the mass ratios of the elements involved in those reactions. The results of his research are known as **Dalton's atomic theory,** which he proposed in 1803. The main points of his theory are summarized in **Table 4.2.** Dalton published his ideas in a book, an extract of which is shown in **Figure 4.2.**

Reading Check **Compare and contrast** Democritus' and Dalton's ideas.



■ **Figure 4.2** In his book *A New System of Chemical Philosophy,* John Dalton presented his symbols for the elements known at that time and their possible combinations.

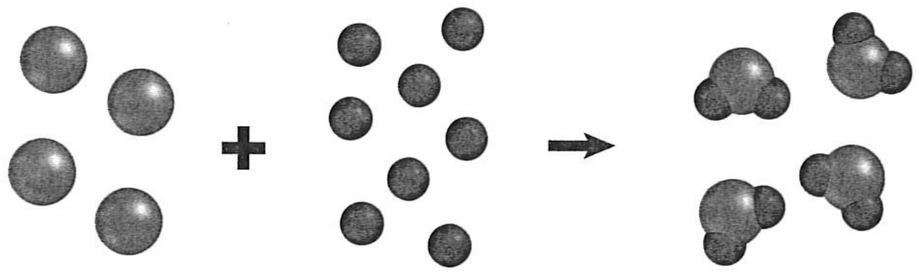


Figure 4.3 When atoms of two or more elements combine to form a compound, the number of atoms of each element is conserved. Thus, the mass is conserved as well.

Conservation of mass Recall from Chapter 3 that the law of conservation of mass states that mass is conserved in any process, such as a chemical reaction. Dalton's atomic theory easily explains that the conservation of mass in chemical reactions is the result of the separation, combination, or rearrangement of atoms—atoms that are not created, destroyed, or divided in the process. The formation of a compound from the combining of elements and the conservation of mass during the process are shown in **Figure 4.3.** The number of atoms of each type is the same before and after the reaction. Dalton's convincing experimental evidence and clear explanation of the composition of compounds, and conservation of mass led to the general acceptance of his atomic theory.

Dalton's atomic theory was a huge step toward the current atomic model of matter. However, not all of Dalton's theory was accurate. As is often the case in science, Dalton's theory had to be revised as additional information was learned that could not be explained by the theory. As you will learn in this chapter, Dalton was wrong about atoms being indivisible. Atoms are divisible into several subatomic particles. Dalton was also wrong about all atoms of a given element having identical properties. Atoms of the same element can have slightly different masses.

Section 4.1 Assessment

Section Summary

- Democritus was the first person to propose the existence of atoms.
- According to Democritus, atoms are solid, homogeneous, and indivisible.
- Aristotle did not believe in the existence of atoms.
- John Dalton's atomic theory is based on numerous scientific experiments.

1. MAIN Idea **Contrast** the methods used by the Greek philosophers and Dalton to study the atom.
2. **Define** *atom* using your own words.
3. **Summarize** Dalton's atomic theory.
4. **Explain** how Dalton's theory of the atom and the conservation of mass are related.
5. **Apply** Six atoms of Element A combine with eight atoms of Element B to produce six compound particles. How many atoms of Elements A and B does each particle contain? Are all of the atoms used to form compounds?
6. **Design** a concept map that compares and contrasts the atomic ideas proposed by Democritus and John Dalton.

Section 4.2

Objectives

- **Define** atom.
- **Distinguish** between the subatomic particles in terms of relative charge and mass.
- **Describe** the structure of the atom, including the locations of the subatomic particles.

Review Vocabulary

model: a visual, verbal, and/or mathematical explanation of data collected from many experiments

New Vocabulary

atom
cathode ray
electron
nucleus
proton
neutron

Defining the Atom

MAIN Idea An atom is made of a nucleus containing protons and neutrons; electrons move around the nucleus.

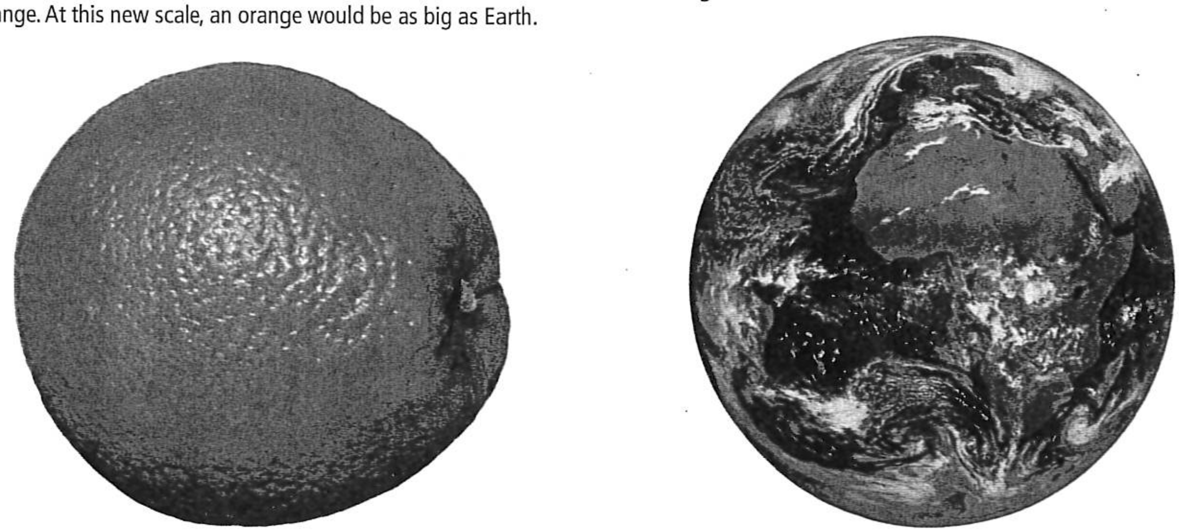
Real-World Reading Link If you have ever accidentally bitten into a peach pit, you know that your teeth pass easily through the fruit, but cannot dent the hard pit. Similarly, many particles that pass through the outer parts of an atom are deflected by the dense center of the atom.

The Atom

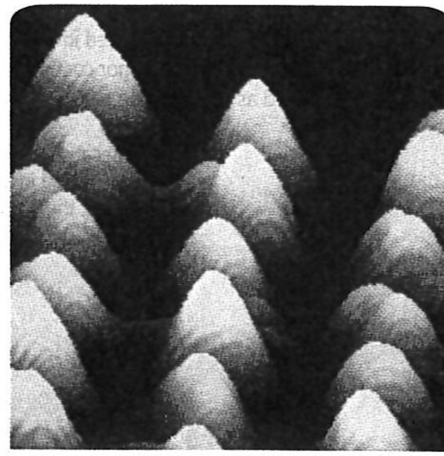
Many experiments since Dalton's time have proven that atoms do exist. So what exactly is the definition of an atom? To answer this question, consider a gold ring. Suppose you decide to grind the ring down into a pile of gold dust. Each fragment of gold dust still retains all of the properties of gold. If it were possible—which it is not without special equipment—you could continue to divide the gold dust particles into still smaller particles. Eventually, you would encounter a particle that could not be divided any further and still retain the properties of gold. This smallest particle of an element that retains the properties of the element is called an **atom.**

To get an idea of its size, consider the population of the world, which was about 6.5×10^9 in 2006. By comparison, a typical solid-copper penny contains 2.9×10^{22} atoms, almost five trillion times the world population! The diameter of a single copper atom is 1.28×10^{-10} m. Placing 6.5×10^9 copper atoms side by side would result in a line of copper atoms less than 1 m long. **Figure 4.4** illustrates another way to visualize the size of an atom. Imagine that you increase the size of an atom to be as big as an orange. To keep the proportions between the real sizes of the atom and of the orange, you would have to increase the size of the orange and make it as big as Earth. This illustrates how small atoms are.

Figure 4.4 Imagine that you could increase the size of an atom to make it as big as an orange. At this new scale, an orange would be as big as Earth.



Connection to Biology **Looking at atoms** You might think that because atoms are so small, there would be no way to see them. However, an instrument called the scanning tunneling microscope (STM) allows individual atoms to be seen. Just as you need a microscope to study cells in biology, the STM allows you to study atoms. An STM works as follows: a fine point is moved above a sample and the interaction of the point with the superficial atoms is recorded electronically. **Figure 4.5** illustrates how individual atoms look when observed with an STM. Scientists are now able to move individual atoms around to form shapes, patterns, and even simple machines. This capability has led to the exciting new field of nanotechnology. The promise of nanotechnology is molecular manufacturing—the atom-by-atom building of machines the size of molecules. As you will read in Chapter 8, a molecule is a group of atoms that are bonded together and act as a unit.



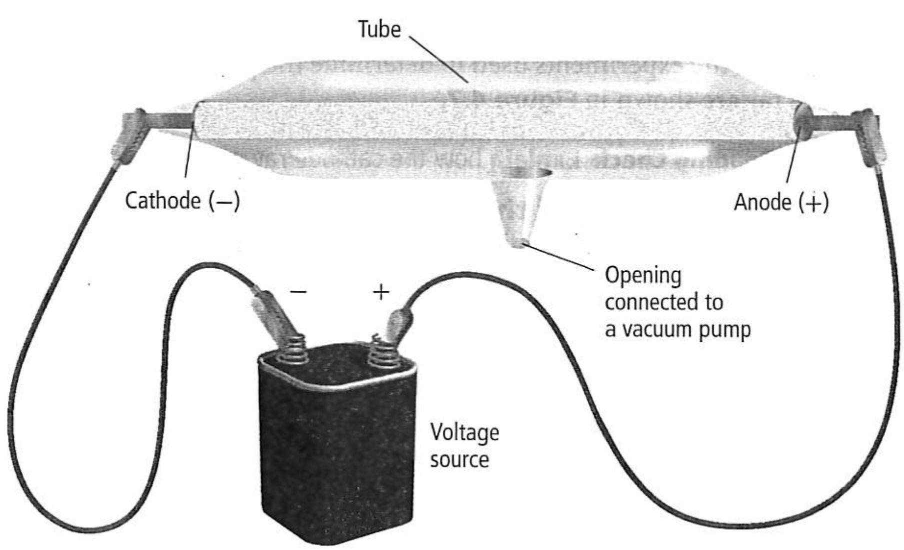
■ **Figure 4.5** This image, recorded with an STM, shows the individual atoms of a fatty acid on a graphite surface. The false colors were added later on to improve the contrast between each atom.

The Electron

Once scientists were convinced of the existence of atoms, a new set of questions emerged. What is an atom like? Is the composition of an atom uniform throughout, or is it composed of still-smaller particles? Although many scientists researched the atom in the 1800s, it was not until almost 1900 that some of these questions were answered.

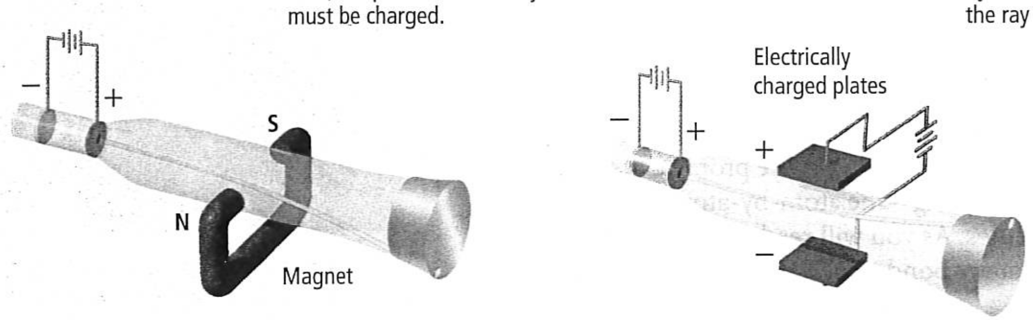
The cathode-ray tube As scientists tried to unravel the atom, they began to make connections between matter and electric charge. For instance, has your hair ever clung to your comb? To explore the connection, some scientists wondered how electricity might behave in the absence of matter. With the help of the newly invented vacuum pump, they passed electricity through glass tubes from which most of the air had been removed. Such tubes are called cathode-ray tubes.

A typical cathode-ray tube used by researchers for studying the relationship between mass and charge is illustrated in **Figure 4.6.** Note that metal electrodes are located at opposite ends of the tube. The electrode connected to the negative terminal of the battery is called the cathode, and the electrode connected to the positive terminal is called the anode.



■ **Figure 4.6** A cathode-ray tube is a tube with an anode at one end and a cathode at the other end. When a voltage is applied, electricity travels from the cathode to the anode.

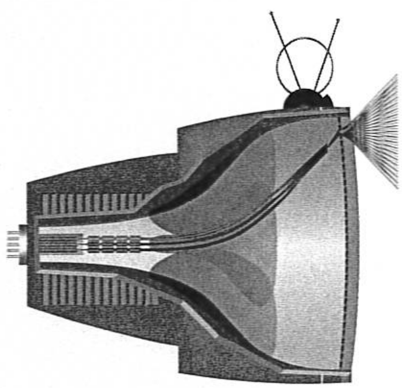
Figure 4.7 A tiny hole located in the center of the anode produces a thin beam of electrons. A phosphor coating allows the position of the beam to be determined as it strikes the end of the tube.



a Because the cathode ray is deflected in a magnetic field, the particles in the ray must be charged.

b Because the cathode ray is deflected toward the positively charged plate by an electric field, the particles in the ray must have a negative charge.

Real-World Chemistry
Cathode Ray



Television Television was invented in the 1920's. Conventional television images are formed as cathode rays strike light-producing chemicals that coat the back of the screen.

FOLDABLES Incorporate information from this section into your Foldable.

Sir William Crookes While working in a darkened laboratory, English physicist Sir William Crookes noticed a flash of light within one of the cathode-ray tubes. A green flash was produced by some form of radiation striking a zinc-sulfide coating that had been applied to the end of the tube. Further work showed that there was a ray (radiation) going through the tube. This ray, originating from the cathode and traveling to the anode, was called a **cathode ray.** The accidental discovery of the cathode ray led to the invention of television. A conventional television is nothing more than a cathode-ray tube.

Scientists continued their research using cathode-ray tubes, and they were fairly convinced by the end of the 1800s of the following:

- Cathode rays were a stream of charged particles.
- The particles carried a negative charge. (The exact value of the negative charge was not known.)

Because changing the metal that makes up the electrodes or varying the gas (at very low pressure) in the cathode-ray tube did not affect the cathode ray produced, researchers concluded that the ray's negative particles were found in all forms of matter. These negatively charged particles that are part of all forms of matter are now known as **electrons.** Some of the experiments used to determine the properties of the cathode ray are shown in **Figure 4.7.**

Reading Check **Explain** how the cathode ray was discovered.

Mass and charge of the electron In spite of the progress made from all of the cathode-ray tube experiments, no one succeeded in determining the mass of a single cathode-ray particle. Unable to measure the particle's mass directly, English physicist J. J. Thomson (1856–1940) began a series of cathode-ray tube experiments at Cambridge University in the late 1890s to determine the ratio of its charge to its mass.

Charge-to-mass ratio By carefully measuring the effects of both magnetic and electric fields on a cathode ray, Thomson was able to determine the charge-to-mass ratio of the charged particle. He then compared that ratio to other known ratios.

Thomson concluded that the mass of the charged particle was much less than that of a hydrogen atom, the lightest known atom. The conclusion was shocking because it meant there were particles smaller than the atom. In other words, Dalton had been incorrect—atoms were divisible into smaller subatomic particles. Because Dalton's atomic theory had become so widely accepted and Thomson's conclusion was so revolutionary, many other scientists found it hard to accept this new discovery. But Thomson was correct. He had identified the first subatomic particle—the electron. He received a Nobel Prize in 1906 for this discovery.

Reading Check **Summarize** how Thomson discovered the electron.

The oil-drop experiment and the charge of an electron

The next significant development came in the early 1910s, when the American physicist Robert Millikan (1868–1953) determined the charge of an electron using the oil-drop apparatus shown in **Figure 4.8.** In this apparatus, oil is sprayed into the chamber above the two parallel charged plates. The top plate has a small hole through which the oil drops. X rays knock out electrons from the air particles between the plates and the electrons stick to the droplets, giving them a negative charge. By varying the intensity of the electric field, Millikan could control the rate of a droplet's fall. He determined that the magnitude of the charge on each drop increased in discrete amounts and determined that the smallest common denominator was 1.602×10^{-19} coulombs. He identified this number as the charge of the electron. This charge was later equated to a single unit of negative charge noted $1-$; in other words, a single electron carries a charge of $1-$.

So good was Millikan's experimental setup and technique that the charge he measured almost one hundred years ago is within 1% of the currently accepted value.

Mass of an electron Knowing the electron's charge and using the known charge-to-mass ratio, Millikan calculated the mass of an electron. The equation below shows how small the mass of an electron is.

$$\text{Mass of an electron} = 9.1 \times 10^{-28}\ \text{g} = \frac{1}{1840}\ \text{the mass of a hydrogen atom}$$

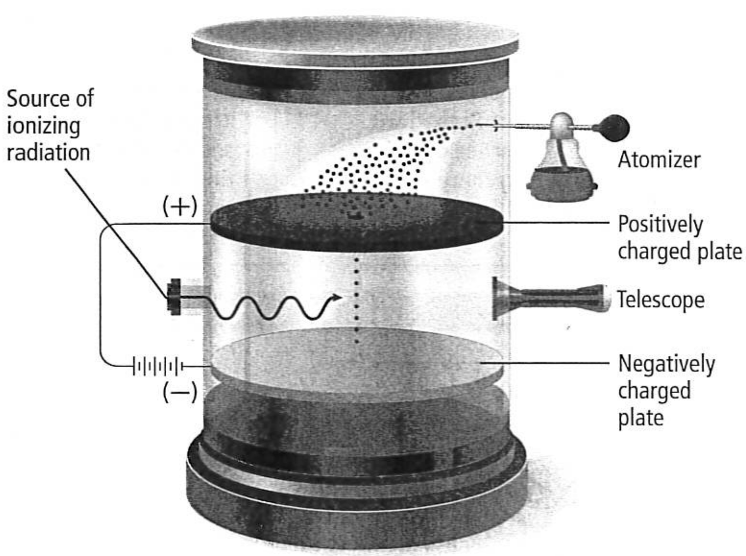
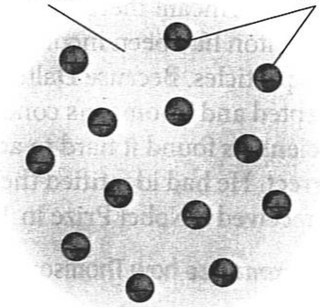


Figure 4.8 The motion of the oil droplets within Millikan's apparatus depends on the charge of droplets and on the electric field. Millikan observed the droplets with the telescope. He could make the droplets fall more slowly, rise, or pause as he varied the strength of the electric field. From his observations, he calculated the charge on each droplet.

■ **Figure 4.9** J. J. Thomson's plum pudding model of the atom states that the atom is a uniform, positively charged sphere containing electrons.



The plum pudding model The existence of the electron and the knowledge of some of its properties raised some interesting new questions about the nature of atoms. It was known that matter is neutral—it has no electric charge. You know that matter is neutral from everyday experience: you do not receive an electric shock (except under certain conditions) when you touch an object. If electrons are part of all matter and they possess a negative charge, how can all matter be neutral? Also, if the mass of an electron is so small, what accounts for the rest of the mass in a typical atom?

In an attempt to answer these questions, J. J. Thomson proposed a model of the atom that became known as the plum pudding model. As you can see in **Figure 4.9,** Thomson's model consisted of a spherically shaped atom composed of a uniformly distributed positive charge in which the individual negatively charged electrons resided. As you are about to read, the plum pudding model of the atom did not last for long. **Figure 4.10** summarizes the numerous steps in understanding the structure of the atom.

Reading Check Explain why Thomson's model was called the plum pudding model.

■ **Figure 4.10**

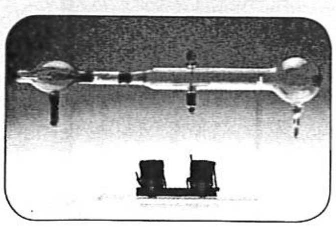
Development of Modern Atomic Theory

Current understanding of the properties and behavior of atoms and subatomic particles is based on the work of scientists worldwide during the past two centuries.

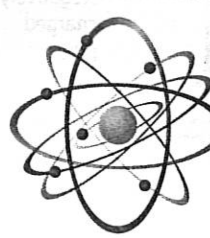
1860 — 1885 — 1910

1911 With the gold foil experiment, Ernest Rutherford determines properties of the nucleus, including charge, relative size, and density.

1932 Scientists develop a particle accelerator to fire protons at lithium nuclei, splitting them into helium nuclei and releasing energy.

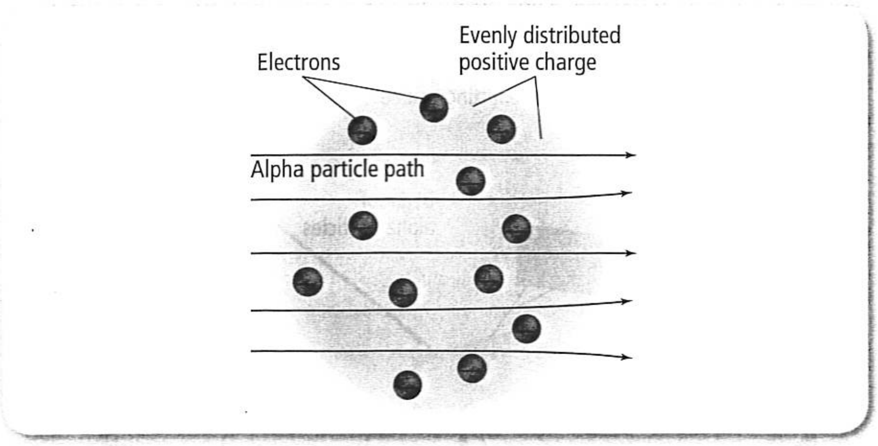


1897 Using cathode-ray tubes, J. J. Thomson identifies the electron and determines the ratio of the mass of an electron to its electric charge.



1913 Niels Bohr publishes a theory of atomic structure relating the electron arrangement in atoms and atomic chemical properties.

1932 James Chadwick proves the existence of neutrons.



■ **Figure 4.11** Based on Thomson's model, Rutherford expected the light alpha particles to pass through gold atoms. He expected only a few of them to be slightly deflected.

The Nucleus

In 1911, Ernest Rutherford (1871–1937) began to study how positively charged alpha particles (radioactive particles you will read more about later in this chapter) interacted with solid matter. With a small group of scientists, Rutherford conducted an experiment to see if alpha particles would be deflected as they passed through a thin gold foil.

Rutherford's experiment In the experiment, a narrow beam of alpha particles was aimed at a thin sheet of gold foil. A zinc-sulfide-coated screen surrounding the gold foil produced a flash of light when struck by an alpha particle. By noting where the flashes occurred, the scientists could determine if the atoms in the gold foil deflected the alpha particles.

Rutherford was aware of Thomson's plum pudding model of the atom. He expected the paths of the massive and fast-moving alpha particles to be only slightly altered by a collision with an electron. And because the positive charge within the gold atoms was thought to be uniformly distributed, he thought it would not alter the paths of the alpha particles, either. **Figure 4.11** shows the results Rutherford expected from the experiment.

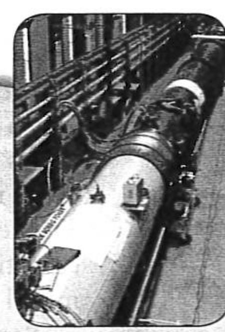
1938 Lise Meitner, Otto Hahn, and Fritz Strassman split uranium atoms in a process they called fission.

1939–1945 Scientists in the United States and Germany each work on projects to develop the first atomic weapon.

1954 CERN, the world's largest nuclear physics research center, located in Switzerland, is founded to study particle physics.

1960

1968 Scientists provide the first experimental evidence for subatomic particles known as quarks.



1985

2007 The Large Hadron Collider at CERN studies the properties of subatomic particles and nuclear matter.

2010

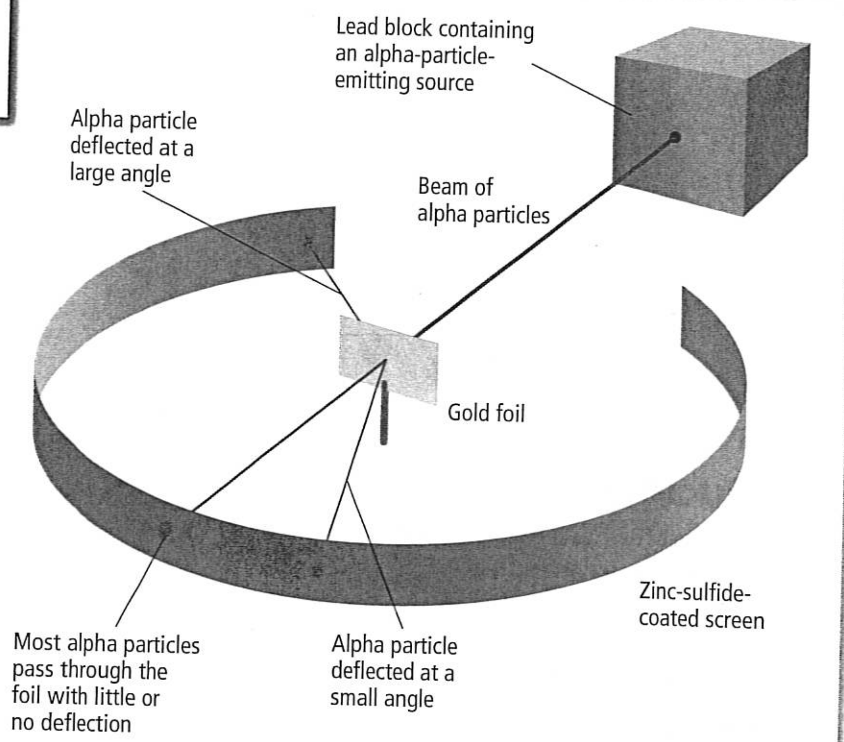
Concepts In Motion

Interactive Time Line To learn more about these discoveries and others, visit glencoe.com.

Chemistry Online

Concepts In Motion

Interactive Figure To see an animation of the gold foil experiment, visit glencoe.com.



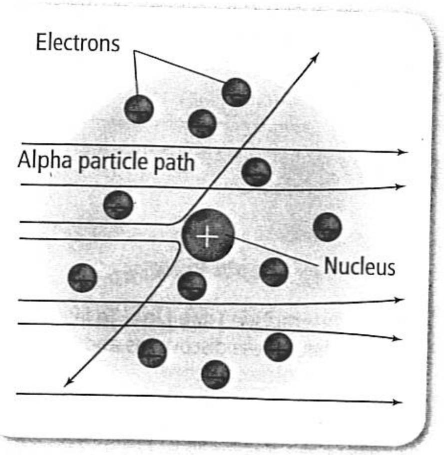
■ **Figure 4.12** During Rutherford's experiment, a beam of alpha particles bombarded a thin gold foil. Most of the alpha particles went through the gold foil. However, a few of them bounced back, some at large angles.

The actual results observed by Rutherford and his colleagues are shown in **Figure 4.12.** A few of the alpha particles were deflected at large angles. Several particles were deflected straight back toward the source. Rutherford likened the results to firing a large artillery shell at a sheet of paper and the shell coming back at the cannon.

Rutherford's model of the atom Rutherford concluded that the plum pudding model was incorrect because it could not explain the results of the gold foil experiment. Considering the properties of the alpha particles and the electrons, and the frequency of the deflections, he calculated that an atom consisted mostly of empty space through which the electrons move. He also concluded that almost all of the atom's positive charge and almost all of its mass were contained in a tiny, dense region in the center of the atom, which he called the **nucleus.** The negatively charged electrons are held within the atom by their attraction to the positively charged nucleus. Rutherford's nuclear atomic model is shown in **Figure 4.13.**

■ **Figure 4.13** In Rutherford's nuclear model, the atom is composed of a dense, positively charged nucleus that is surrounded by negative electrons. Alpha particles passing far from the nucleus are only slightly deflected. Alpha particles directly approaching the nucleus are deflected at large angles.

Infer *what force causes the deflection of alpha particles.*



Because the nucleus occupies such a small space and contains most of an atom's mass, it is incredibly dense. If a nucleus were the size of the dot in the exclamation point at the end of this sentence, its mass would be approximately as much as that of 70 automobiles! The volume of space through which the electrons move is huge compared to the volume of the nucleus. A typical atom's diameter is approximately 10,000 times the diameter of the nucleus. If an atom had a diameter of two football fields, the nucleus would be the size of a nickel.

Reading Check **Describe** Rutherford's model of the atom.

The repulsive force produced between the positive nucleus and the positive alpha particles causes the deflections. **Figure 4.13** illustrates how Rutherford's nuclear atomic model explained the results of the gold foil experiment. The nuclear model also explains the neutral nature of matter: the positive charge of the nucleus balances the negative charge of the electrons. However, the model still could not account for all of the atom's mass.

The proton and the neutron By 1920, Rutherford had refined the concept of the nucleus and concluded that the nucleus contained positively charged particles called protons. A **proton** is a subatomic particle carrying a charge equal to but opposite that of an electron; that is, a proton has a charge of 1+. In 1932, Rutherford's coworker, English physicist James Chadwick (1891–1974), showed that the nucleus also contained another subatomic neutral particle, called the neutron. A **neutron** is a subatomic particle that has a mass nearly equal to that of a proton, but it carries no electric charge. In 1935, Chadwick received the Nobel Prize in Physics for proving the existence of neutrons.

VOCABULARY

SCIENCE USAGE V. COMMON USAGE

Neutral

Science usage: to have no electric charge

Neutrons have a charge of zero. They are neutral particles.

Common usage: not engaged in either side

Switzerland remained neutral during World War II.

DATA ANALYSIS LAB

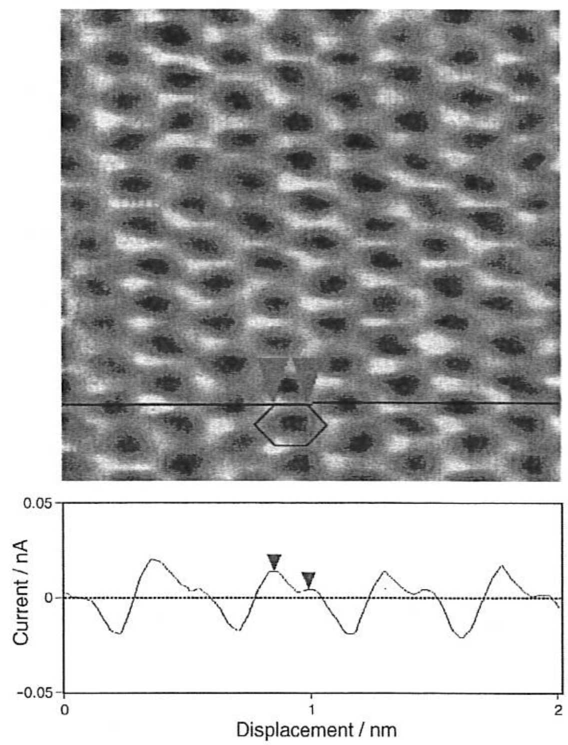
Based on Real Data*

Interpret Scientific Illustrations

What are the apparent atomic distances of carbon atoms in a well-defined crystalline material? To visualize individual atoms, a group of scientists used a scanning tunneling microscope (STM) to test a crystalline material called highly ordered pyrolytic graphite (HOPG). An STM is an instrument used to perform surface atomic-scale imaging.

Data and Observations

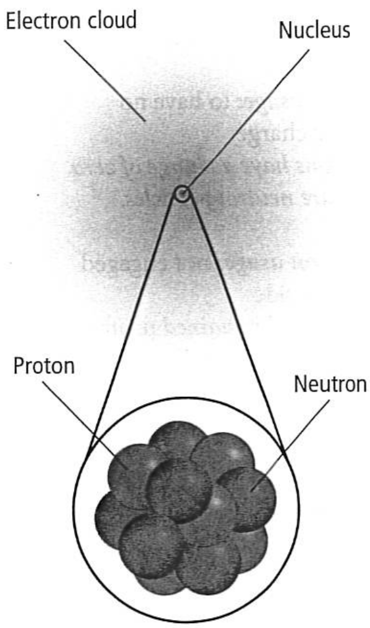
The image shows all of the carbon atoms in the surface layer of the graphite material. Each hexagonal ring, indicated by the drawing in the figure, consists of three brighter spots separated by three fainter spots. These bright spots are from alternate carbon atoms in the surface layer of the graphite structure. The cross-sectional view below the photo corresponds to the line drawn in the image. It indicates the atomic periodicity and apparent atomic distances.



*Data obtained from: Chaun-Jian Zhong et al. 2003. Atomic scale imaging: a hands-on scanning probe microscopy laboratory for undergraduates. *Journal of Chemical Education* 80: 194–197.

Think Critically

1. **Estimate** the distance between two nearest bright spots.
2. **Estimate** the distance between two nearest neighbor spots (brighter–fainter, marked with triangles in the figure).
3. **State** What do the black spots in the image represent?
4. **Explain** How many carbon atoms are across the line drawn in the image?



■ **Figure 4.14** Atoms are composed of a a nucleus containing protons and neutrons, and surrounded by a cloud of electrons.

Concepts In Motion

Interactive Figure To see an animation of the structure of the atom, visit glencoe.com.

Table 4.3 Properties of Subatomic Particles

Concepts In Motion

Interactive Table Explore the properties of subatomic particles at glencoe.com.

Particle	Symbol	Location	Relative Electric Charge	Relative Mass	Actual Mass (g)
Electron	e^-	In the space surrounding the nucleus	1–	$\frac{1}{1840}$	9.11×10^{-28}
Proton	p	In the nucleus	1+	1	1.673×10^{-24}
Neutron	n	In the nucleus	0	1	1.675×10^{-24}

Completing the model of the atom All atoms are made up of the three fundamental subatomic particles—the electron, the proton, and the neutron. Atoms are spherically shaped, with a small, dense nucleus of positive charge surrounded by one or more negatively charged electrons. Most of an atom consists of fast-moving electrons traveling through the empty space surrounding the nucleus. The electrons are held within the atom by their attraction to the positively charged nucleus. The nucleus, which is composed of neutral neutrons (hydrogen's single-proton nucleus is an exception) and positively charged protons, contains all of an atom's positive charge and more than 99.97% of its mass. It occupies only about one ten-thousandth of the volume of the atom. Because an atom is electrically neutral, the number of protons in the nucleus equals the number of electrons surrounding the nucleus. The features of a typical atom are shown in **Figure 4.14,** and the properties of the fundamental subatomic particles are summarized in **Table 4.3.**

Subatomic particle research is still a major interest to modern scientists. In fact, scientists have determined that protons and neutrons have their own structures. They are composed of subatomic particles called quarks. These particles will not be covered in this textbook because scientists do not yet understand if or how they affect chemical behavior. As you will learn in later chapters, chemical behavior can be explained by considering only an atom's electrons.

Section 4.2 Assessment

Section Summary

- An atom is the smallest particle of an element that maintains the properties of that element.
- Electrons have a 1– charge, protons have a 1+ charge, and neutrons have no charge.
- An atom consists mostly of empty space surrounding the nucleus.

7. MAIN Idea **Describe** the structure of a typical atom. Identify where each subatomic particle is located.
8. **Compare and contrast** Thomson's plum pudding atomic model with Rutherford's nuclear atomic model.
9. **Evaluate** the experiments that led to the conclusion that electrons are negatively charged particles found in all matter.
10. **Compare** the relative charge and mass of each of the subatomic particles.
11. **Calculate** What is the difference expressed in kilograms between the mass of a proton and the mass of an electron?

ChemistryOnline Self-Check Quiz glencoe.com

Section 4.3

Objectives

- **Explain** the role of atomic number in determining the identity of an atom.
- **Define** an isotope.
- **Explain** why atomic masses are not whole numbers.
- **Calculate** the number of electrons, protons, and neutrons in an atom given its mass number and atomic number.

Review Vocabulary

periodic table: a chart that organizes all known elements into a grid of horizontal rows (periods) and vertical columns (groups or families) arranged by increasing atomic number

New Vocabulary

atomic number
isotope
mass number
atomic mass unit (amu)
atomic mass

How Atoms Differ

MAIN Idea The number of protons and the mass number define the type of atom.

Real-World Reading Link You are probably aware that numbers are used every day to identify people and objects. For example, people can be identified by their Social Security numbers and computers by their IP addresses. Atoms and nuclei are also identified by numbers.

Atomic Number

As shown in the periodic table of the elements inside the back cover of this textbook, there are more than 110 different elements. What makes an atom of one element different from an atom of another element?

Not long after Rutherford's gold foil experiment, the English scientist Henry Moseley (1887–1915) discovered that atoms of each element contain a unique positive charge in their nuclei. Thus, the number of protons in an atom identifies it as an atom of a particular element. The number of protons in an atom is referred to as the **atomic number.** The information provided by the periodic table for hydrogen is shown in **Figure 4.15.** The number 1 above the symbol for hydrogen (H) is the number of protons, or the atomic number. Moving across the periodic table to the right, you will next come to helium (He). It has two protons in its nucleus, and thus it has an atomic number of 2. The next row begins with lithium (Li), atomic number 3, followed by beryllium (Be), atomic number 4, and so on. The periodic table is organized left-to-right and top-to-bottom by increasing atomic number.

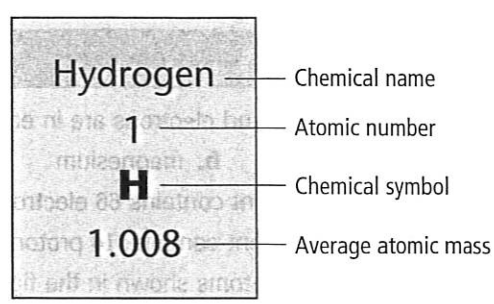
Because all atoms are neutral, the number of protons and electrons in an atom must be equal. Thus, once you know the atomic number of an element, you know the number of protons and the number of electrons an atom of that element contains. For example, an atom of lithium, atomic number 3, contains three protons and three electrons.

Atomic number

$$\text{atomic number} = \text{number of protons} = \text{number of electrons}$$

The atomic number of an atom equals its number of protons and its number of electrons.

Figure 4.15 In the periodic table, each element is represented by its chemical name, atomic number, chemical symbol, and average atomic mass. **Determine** *the number of protons and the number of electrons in an atom of gold.*



EXAMPLE Problem 4.1

Math Handbook
Solving Algebraic Equations
pages 954–955

Atomic Number Complete the following table.

Composition of Several Elements

	Element	Atomic Number	Protons	Electrons
a.	Pb	82		
b.			8	
c.				30

1 Analyze the Problem

Apply the relationship among atomic number, number of protons, and number of electrons to complete most of the table. Then, use the periodic table to identify the element.

Known

a. element = Pb, atomic number = 82
b. number of protons = 8
c. number of electrons = 30

Unknown

a. number of protons **(N_p)**, number of electrons **(N_e) = ?**
b. element, atomic number **(Z), N_e = ?**
c. element, Z, N_p = ?

2 Solve for the Unknown

a. number of protons = atomic number — **Apply the atomic-number relationship.**

$N_p = 82$ — **Substitute atomic number = 82.**

number of electrons = number of protons

$N_e = 82$

The number of protons and the number of electrons is 82.

b. atomic number = number of protons — **Apply the atomic-number relationship.**

$Z = 8$ — **Substitute number of protons = 8.**

number of electrons = number of protons

$N_e = 8$

The atomic number and the number of electrons is 8.

The **element** is **oxygen (O).** — **Consult the periodic table to identify the element.**

c. number of protons = number of electrons — **Apply the atomic-number relationship.**

$N_p = 30$ — **Substitute number of electrons = 30.**

atomic number = number of protons

$Z = 30$

The atomic number and the number of protons is 30.

The **element** is **zinc (Zn).** — **Consult the periodic table to identify the element.**

3 Evaluate the Answer

The answers agree with atomic numbers and element symbols given in the periodic table.

PRACTICE Problems

Extra Practice Pages 977–978 and glencoe.com

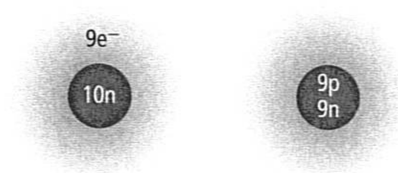
12. How many protons and electrons are in each atom?

a. radon **b.** magnesium

13. An atom of an element contains 66 electrons. Which element is it?

14. An atom of an element contains 14 protons. Which element is it?

15. Challenge Do the atoms shown in the figure to the right have the same atomic number?



Isotopes and Mass Number

Dalton was incorrect about atoms being indivisible and in stating that all atoms of an element are identical. All atoms of an element have the same number of protons and electrons, but the number of neutrons might differ. For example, there are three types of potassium atoms that occur naturally. All three types contain 19 protons and 19 electrons. However, one type of potassium atom contains 20 neutrons, another 21 neutrons, and still another 22 neutrons. Atoms with the same number of protons but different numbers of neutrons are called **isotopes.**

Mass of isotopes Isotopes containing more neutrons have a greater mass. In spite of these differences, isotopes of an atom have the same chemical behavior. As you will read later in this textbook, chemical behavior is determined only by the number of electrons an atom has.

Isotope notation Each isotope of an element is identified with a number called the mass number. The **mass number** is the sum of the atomic number (or number of protons) and neutrons in the nucleus.

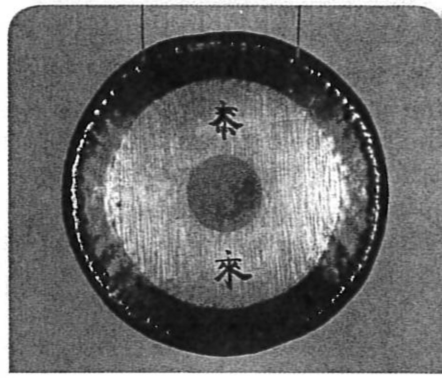
Mass number

mass number = atomic number + number of neutrons

The mass number of an atom is the sum of its atomic number and its number of neutrons.

For example, copper has two isotopes. The isotope with 29 protons and 34 neutrons has a mass number of 63 ($29 + 34 = 63$), and is called copper-63 (also written ^{63}Cu or Cu-63). The isotope with 29 protons and 36 neutrons is called copper-65. Chemists often write out isotopes using a notation involving the chemical symbol, atomic number, and mass number, as shown in **Figure 4.16.**

Natural abundance of isotopes In nature, most elements are found as mixtures of isotopes. Usually, no matter where a sample of an element is obtained, the relative abundance of each isotope is constant. For example, in a banana, 93.26% of the potassium atoms have 20 neutrons, 6.73% have 22 neutrons, and 0.01% have 21 neutrons. In another banana, or in a different source of potassium, the percentage composition of the potassium isotopes will still be the same. The three potassium isotopes are summarized in **Figure 4.17.**



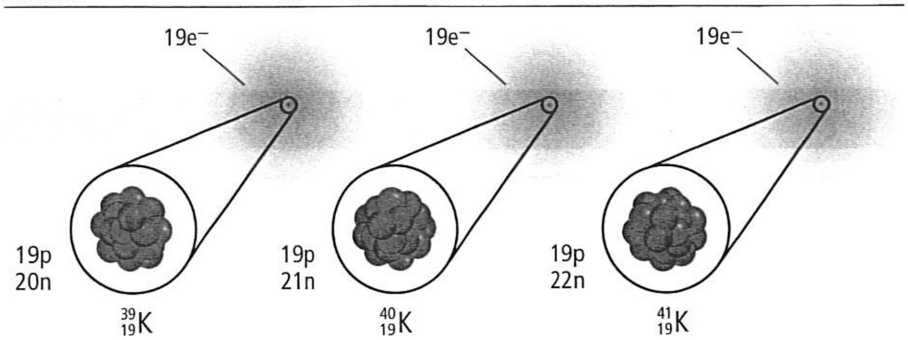
Mass number

$^{63}_{29}Cu$ $^{65}_{29}Cu$

Atomic number

■ **Figure 4.16** Cu is the chemical symbol for copper. Copper, which was used to make this Chinese gong, is composed of 69.2% copper-63 and 30.8% copper- 65.

	Potassium-39	Potassium-40	Potassium-41
Protons	19	19	19
Neutrons	20	21	22
Electrons	19	19	19



■ **Figure 4.17** Potassium has three naturally occurring isotopes: potassium-39, potassium-40, and potassium-41.

List ***the number of protons, neutrons, and electrons in each potassium isotope.***

EXAMPLE Problem 4.2

Use Atomic Number and Mass Number A chemistry laboratory has analyzed the composition of isotopes of several elements. The composition data is given in the table below. Determine the number of protons, electrons, and neutrons in the isotope of neon. Name the isotope and give its symbol.

Isotope Composition Data

	Element	Atomic Number	Mass Number
a.	Neon	10	22
b.	Calcium	20	46
c.	Oxygen	8	17
d.	Iron	26	57
e.	Zinc	30	64
f.	Mercury	80	204

1 Analyze the Problem

You are given some data for neon in the table. The symbol for neon can be found on the periodic table. From the atomic number, the number of protons and electrons in the isotope are known. The number of neutrons in the isotope can be found by subtracting the atomic number from the mass number.

Known
element: neon
atomic number = 10
mass number = 22

Unknown
number of protons **(N_p)**, electrons **(N_e)**, and neutrons **(N_n)** = **?**
name of isotope = **?**
symbol for isotope = **?**

2 Solve for the Unknown

number of protons = atomic number = **10** — **Apply the atomic number relationship.**

number of electrons = atomic number = **10**

number of neutrons = mass number − atomic number — **Use the atomic number and the mass number to calculate the number of neutrons.**

$N_n = 22 - 10 = \mathbf{12}$ — **Substitute mass number = 22 and atomic number = 10**

The **name** of the isotope is **neon-22.** — **Use the element name and mass number to write the isotope's name.**

The **symbol** for the isotope is $^{22}_{10}\mathbf{Ne}$. — **Use the chemical symbol, mass number, and atomic number to write out the isotope in symbolic notation form.**

3 Evaluate the Answer

The relationships among number of electrons, protons, and neutrons have been applied correctly. The isotope's name and symbol are in the correct format. Refer to pages 944–945 the Elements Handbook to learn more about neon.

PRACTICE Problems

Extra Practice Page 978 and glencoe.com

16. Determine the number of protons, electrons, and neutrons for isotopes **b.–f.** in the table above. Name each isotope, and write its symbol.

17. Challenge An atom has a mass number of 55. Its number of neutrons is the sum of its atomic number and five. How many protons, neutrons, and electrons does this atom have? What is the identity of this atom?

Table 4.4	Masses of Subatomic Particles
Particle	**Mass (amu)**
Electron	0.000549
Proton	1.007276
Neutron	1.008665

Mass of Atoms

Recall from **Table 4.3** that the masses of both protons and neutrons are approximately 1.67×10^{-24} g. While this is a small mass, the mass of an electron is even smaller—only about 1/1840 that of a proton or a neutron.

Atomic mass unit Because these extremely small masses expressed in scientific notation are difficult to work with, chemists have developed a method of measuring the mass of an atom relative to the mass of a specific atomic standard. That standard is the carbon-12 atom. Scientists assigned the carbon-12 atom a mass of exactly 12 atomic mass units. Thus, one **atomic mass unit (amu)** is defined as one-twelfth the mass of a carbon-12 atom. Although a mass of 1 amu is nearly equal to the mass of a single proton or a single neutron, it is important to realize that the values are slightly different. **Table 4.4** gives the masses of the subatomic particles in terms of amu.

Atomic mass Because an atom's mass depends mainly on the number of protons and neutrons it contains, and because protons and neutrons have masses close to 1 amu, you might expect the atomic mass of an element to always be nearly a whole number. However, this is often not the case. The explanation involves how atomic mass is defined. The **atomic mass** of an element is the weighted average mass of the isotopes of that element. Because isotopes have different masses, the weighted average is not a whole number. The calculation of the atomic mass of chlorine is illustrated in **Figure 4.18.**

VOCABULARY

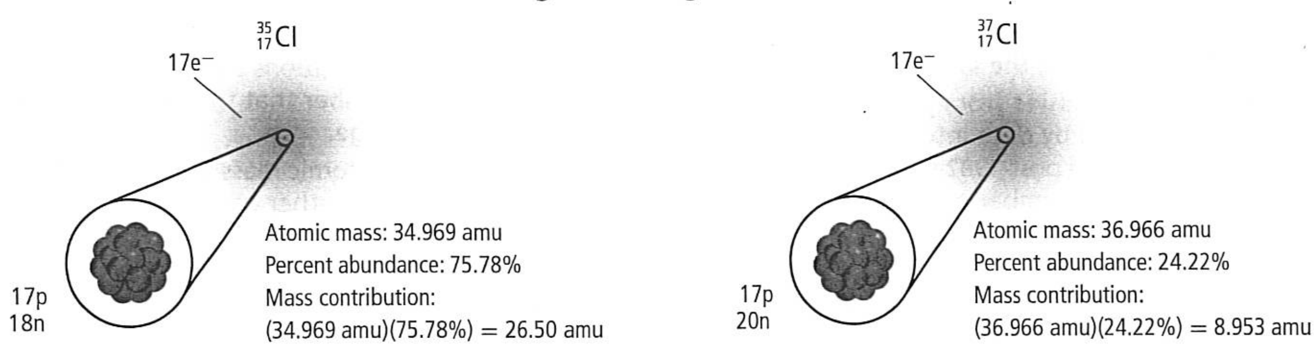
ACADEMIC VOCABULARY

Specific
characterized by precise formulation or accurate restriction
Some diseases have specific symptoms.

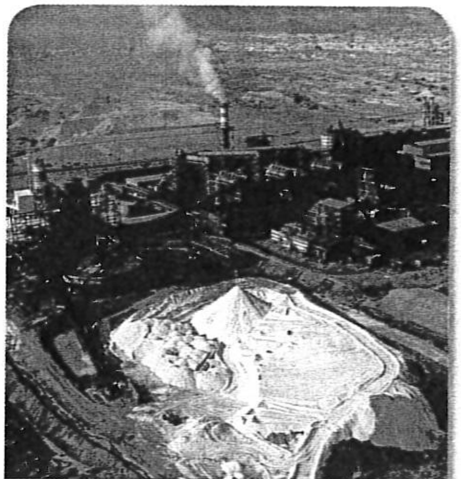
ChemistryOnline
Personal Tutor For an online tutorial on finding an average, visit glencoe.com.

■ **Figure 4.18** To calculate the weighted average atomic mass of chlorine, you first need to calculate the mass contribution of each isotope.

Calculate the Weighted Average Atomic Mass of Chlorine



Weighted average atomic mass of chlorine = (26.50 amu + 8.953 amu) = 35.45 amu



■ **Figure 4.19** Bromine is extracted from sea water and salt lakes. The Dead Sea area in Israel is one of the major bromine production sites in the world. Applications of bromine include microbe and algae control in swimming pools and flame-retardants. It is also used in medicines, oils, paints, and pesticides.

Chlorine exists naturally as a mixture of about 76% chlorine-35 and 24% chlorine-37. It has an atomic mass of 35.453 amu. Because atomic mass is a weighted average, the chlorine-35 atoms, which exist in greater abundance than the chlorine-37 atoms, have a greater effect in determining the atomic mass. The atomic mass of chlorine is calculated by multiplying each isotope's percent abundance by its atomic mass and then adding the products. The process is similar to calculating an average grade. You can calculate the atomic mass of any element if you know the number of naturally occurring isotopes, their masses, and their percent abundances.

Reading Check **Explain** how to calculate atomic mass.

Isotope abundances Analyzing an element's mass can indicate the most abundant isotope for that element. For example, fluorine (F) has an atomic mass that is extremely close to 19 amu. If fluorine had several fairly abundant isotopes, its atomic mass would not likely be so close to a whole number. Thus, you might conclude that all naturally occurring fluorine is probably in the form of fluorine-19 ($^{19}_{9}F$). Indeed, 100% of naturally occurring fluorine is in the form of fluorine-19. While this type of reasoning generally works well, it is not foolproof. Consider bromine (Br). It has an atomic mass of 79.904 amu. With a mass so close to 80 amu, it seems likely that the most common bromine isotope would be bromine-80. However, Bromine's two isotopes are bromine-79 (78.918 amu, 50.69%) and bromine-81 (80.917 amu, 49.31%). There is no bromine-80 isotope. **Figure 4.19** shows one of the major production sites of bromine, located in the Dead Sea area. Refer to page 940 of the Elements Handbook to learn more about chlorine, fluorine, and bromine.

MiniLab

Model Isotopes

How can you calculate the atomic mass of an element using the percentage abundance of its isotopes? Because they have different compositions, pre- and post-1982 pennies can be used to model an element with two naturally occurring isotopes. From the penny 'isotope' data, you can determine the mass of each penny isotope and the average mass of a penny.

Procedure

1. Read and complete the lab safety form.
2. Get a bag of **pennies** from your teacher, and sort the pennies by date into two groups: pre-1982 pennies and post-1982 pennies. Count and record the total number of pennies and the number in each group.
3. Using a **balance**, determine the mass of 10 pennies from each group. Record each mass to the nearest 0.01 g. Divide the total mass of each group by 10 to get the average mass of a pre- and post-1982 penny isotope.

Analysis

1. **Calculate** the percentage abundance of each group using data from Step 2. To do this, divide the number of pennies in each group by the total number of pennies.
2. **Determine** the atomic mass of a penny using the percentage abundance of each "isotope" and data from Step 3. To do this, use the following equation:

   mass contribution = (% abundance)(mass)

   Total the mass contributions to determine the atomic mass. Remember that the percent abundance is a percentage.
3. **Infer** whether the atomic mass would be different if you received another bag of pennies containing a different mixture of pre- and post-1982 pennies. Explain your reasoning.
4. **Explain** why the average mass of each type of penny was determined by measuring 10 pennies instead of by measuring and using the mass of a single penny from each group.