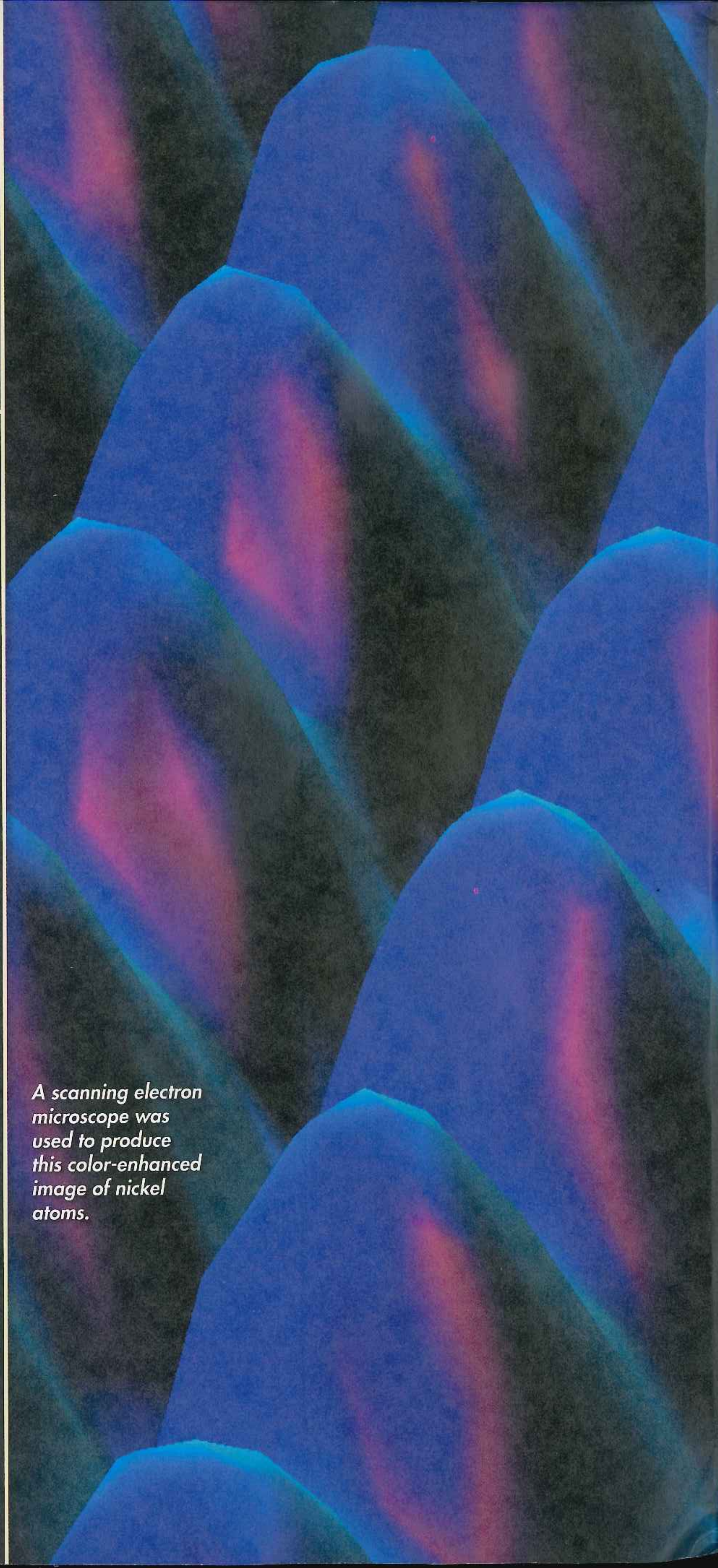
4

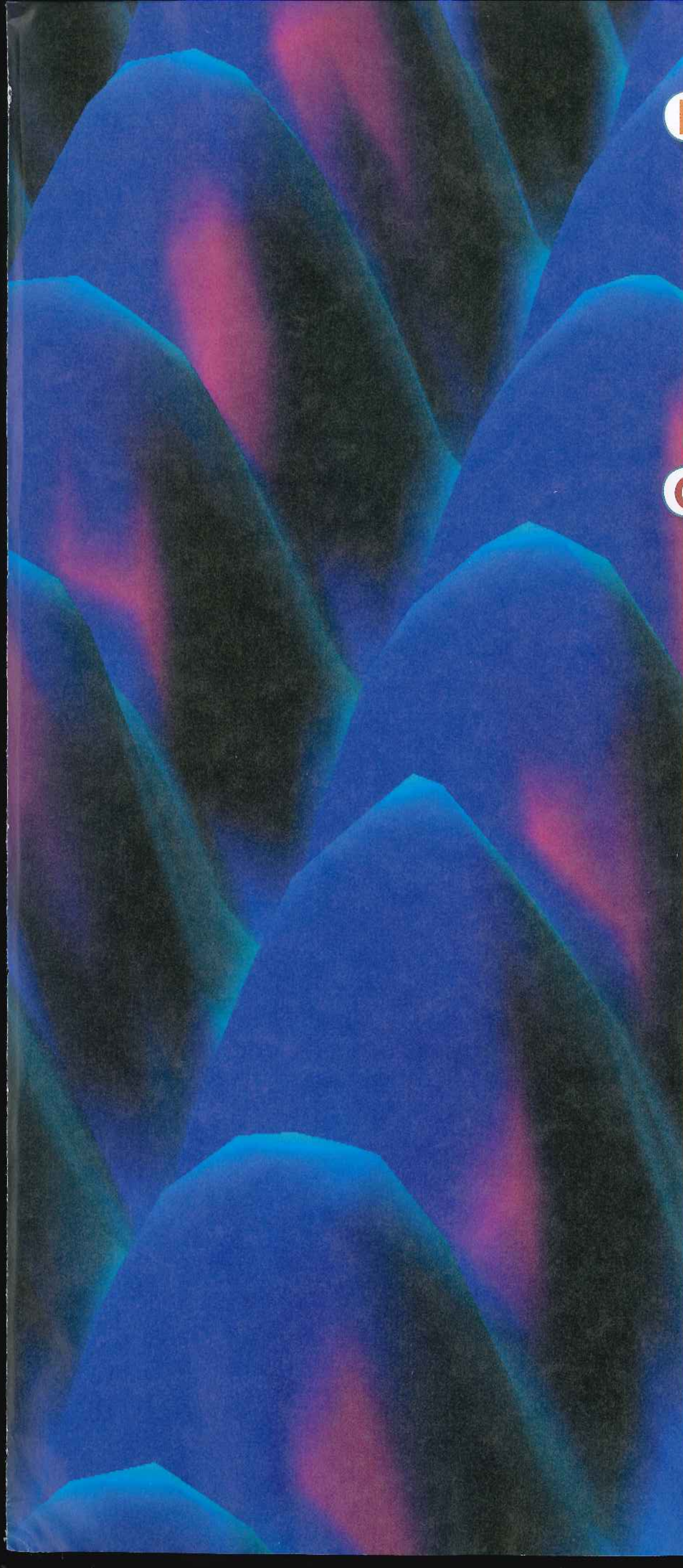
Atomic Structure

INSIDE:

PearsonChem.com



A scanning electron microscope was used to produce this color-enhanced image of nickel atoms.



BIGIDEA

ELECTRONS AND THE STRUCTURE OF ATOMS

Essential Questions:

1. *What components make up an atom?*
2. *How are atoms of one element different from atoms of another element?*

CHEMYSTERY

Artifact or Artifake?

Crystal skulls are shaped like a human skull and carved from quartz crystal. Crystal skulls are thought to have originated from pre-Columbian Central American cultures. If so, then crystal skulls would have been carved several hundred or even thousands of years ago. They would probably have been carved using primitive stone, wooden, and bone tools.

Although crystal skulls are displayed in museums throughout the world, none of them were found in an actual archaeological dig. This unusual circumstance has led to some debate about the history of the skulls. People have questioned whether crystal skulls were ever carved by people from ancient civilizations. Are these sculptures true artifacts that were carved in the pre-Columbian era, or are they just fakes?

▶ Connect to the **BIG**IDEA As you read about the structure of atoms, think about how scientists could identify whether a crystal skull is from an ancient civilization or is just a fake.

NATIONAL SCIENCE EDUCATION STANDARDS

B-1, E-1, G-1

4.1 Defining the Atom

CHEMISTRY & YOU

Q: *How do you study something that you cannot see?* It is sometimes fun to try to figure out what is inside a present before opening it. You could look at the shape or weight of the box. Or maybe you would shake the box a little to find out if anything moved around or made noise inside the box. Similar to how you might study a giftwrapped present, scientists often study things that cannot be seen with the naked eye. In this lesson, you will learn how scientists obtained information about the atoms that they couldn't see.

Key Questions

How did the concept of the atom change from the time of Democritus to the time of John Dalton?

What instruments are used to observe individual atoms?

Vocabulary

- atom
- Dalton's atomic theory

READING SUPPORT

Build Vocabulary: *Word Origins* *Atom* comes from the Greek word *atomos,* meaning "indivisible." ***How does the word origin of atom relate to Dalton's atomic theory?***

Early Models of the Atom

How did the concept of the atom change from the time of Democritus to the time of John Dalton?

Using your unaided eyes, you cannot see the tiny fundamental particles that make up matter. Yet, all matter is composed of such particles, which are called atoms. An **atom** is the smallest particle of an element that retains its identity in a chemical reaction.

The concept of the atom intrigued a number of early scholars. Although these philosophers and scientists could not observe individual atoms, they still were able to propose ideas about the structure of atoms.

Democritus's Atomic Philosophy The Greek philosopher Democritus (460 B.C.–370 B.C.) was among the first to suggest the existence of atoms. **Democritus reasoned that atoms were indivisible and indestructible.** Although Democritus's ideas agreed with later scientific theory, they did not explain chemical behavior. They also lacked experimental support because Democritus's approach was not based on the scientific method.

Dalton's Atomic Theory The real nature of atoms and the connection between observable changes and events at the atomic level were not established for more than 2000 years after Democritus's death. The modern process of discovery regarding atoms began with John Dalton (1766–1844), an English chemist and schoolteacher. **By using experimental methods, Dalton transformed Democritus's ideas on atoms into a scientific theory.** Dalton studied the ratios in which elements combine in chemical reactions.

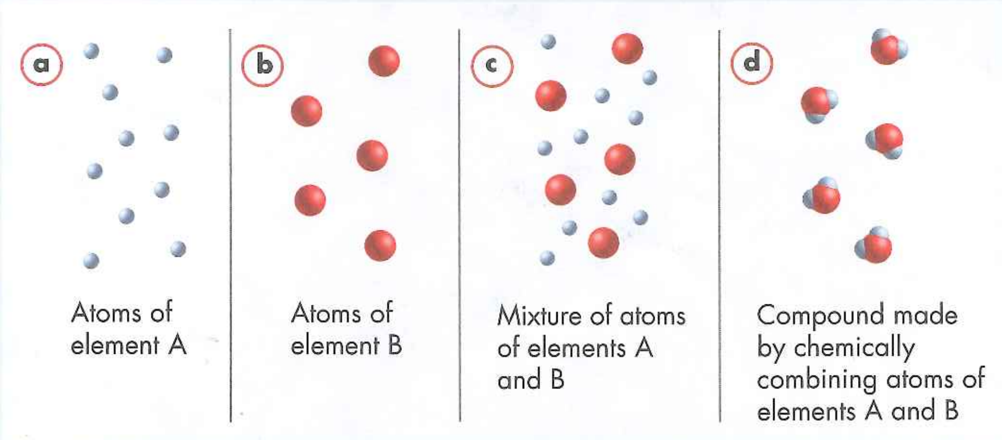


Figure 4.1 Dalton's Atomic Theory
According to Dalton's atomic theory, an element is composed of only one kind of atom, and a compound is composed of particles that are chemical combinations of different kinds of atoms.
Interpret Diagrams ***How does a mixture of atoms of different elements differ from a compound?***

Based on the results of his experiments, Dalton formulated hypotheses and theories to explain his observations. The result of his work is known as **Dalton's atomic theory,** which includes the ideas illustrated in Figure 4.1 and listed below.

1. All elements are composed of tiny indivisible particles called atoms.
2. Atoms of the same element are identical. The atoms of any one element are different from those of any other element.
3. Atoms of different elements can physically mix together or can chemically combine in simple whole-number ratios to form compounds.
4. Chemical reactions occur when atoms are separated from each other, joined, or rearranged in a different combination. Atoms of one element, however, are never changed into atoms of another element as a result of a chemical reaction.

CHEMISTRY & YOU

Q: *How was John Dalton able to study atoms even though he couldn't observe them directly? What evidence did he use to formulate his atomic theory?*

Sizing up the Atom

What instruments are used to observe individual atoms?

The liquid mercury in Figure 4.2 illustrates Dalton's concept of the atom. Whether the size of the drop of mercury is large or small, all drops have the same properties because they are all made of the same kind of atoms.

A coin the size of a penny and composed of pure copper (Cu) is another example. If you were to grind the copper coin into a fine dust, each speck in the small pile of shiny red dust would still have the properties of copper. If by some means you could continue to make the copper dust particles smaller, you would eventually come upon a particle of copper that could no longer be divided and still have the chemical properties of copper. This final particle is an atom.

Atoms are very small. A pure copper coin the size of a penny contains about 2×10^{22} atoms. By comparison, Earth's population is only about 7×10^{9} people. There are about 3×10^{12} times as many atoms in the coin as there are people on Earth. If you could line up 100,000,000 copper atoms side by side, they would produce a line only 1 cm long!

Figure 4.2 Drops of Mercury
This petri dish contains drops of liquid mercury. Every drop, no matter its size, has the same properties. Even if you could make a drop the size of one atom, it would still have the chemical properties of mercury.

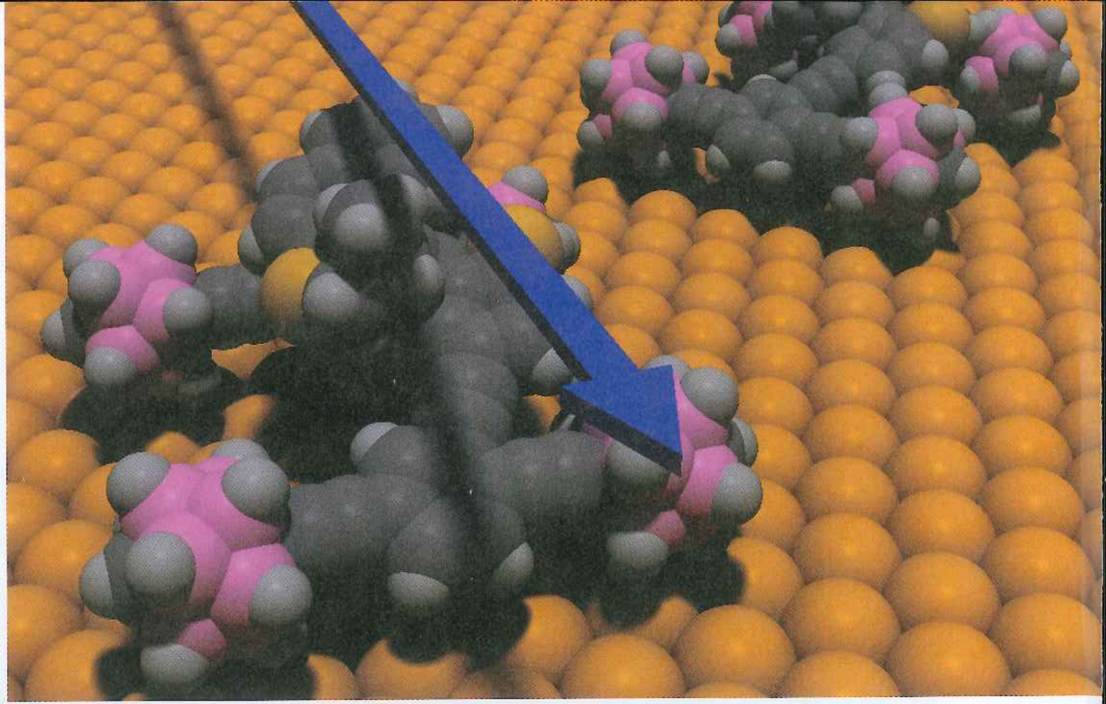


Figure 4.3 Model of a Nanocar
These nanocars are each made of a single molecule. Each nanocar is only about 2 nanometers across. A light-activated paddle wheel on the car propels the car so it can move. The arrow represesents the direction the nanocar moves.

Learn more about the size of the atom *online.*

The radii of most atoms fall within the range of 5×10^{-11} m to 2×10^{-10} m. Does seeing individual atoms seem impossible? **Despite their small size, individual atoms are observable with instruments such as scanning electron microscopes.** In scanning electron microscopes, a beam of electrons is focused on the sample. Electron microscopes are capable of much higher magnifications than light microscopes.

With the help of electron microscopes, individual atoms can even be moved around and arranged in patterns. The ability to move individual atoms holds future promise for the creation of atomic-sized electronic devices, such as circuits and computer chips. An example of a device made from individual atoms is the nanocar shown in Figure 4.3. This atomic-scale, or "nanoscale," technology could become essential to future applications in medicine, communications, solar energy, and space exploration.

4.1 LessonCheck

ONLINE PROBLEMS

1. **Review** How did Democritus characterize atoms?
2. **Explain** How did Dalton advance the atomic philosophy proposed by Democritus?
3. **Identify** What instrument can be used to observe individual atoms?
4. **Explain** In your own words, explain the main ideas of Dalton's atomic theory.
5. **Evaluate** Explain why the ideas on atoms proposed by Dalton constitute a theory, while the ideas proposed by Democritus do not.
6. **Identify** What is the range of the radii of most atoms in nanometers (nm)?
7. **Calculate** A sample of copper with a mass of 63.5 g contains 6.02×10^{23} atoms. Calculate the mass of a single copper atom.

BIGIDEA
ELECTRONS AND THE STRUCTURE OF ATOMS

8. According to Dalton's theory, is it possible to convert atoms of one element into atoms of another? Explain.

4.2 Structure of the Nuclear Atom

CHEMISTRY & YOU

Q: *You can X-ray a person's hand to see inside it—but how can you see inside an atom?* You may have seen X-rays like the one of the hand shown here. Doctors often use X-rays to see bones and other structures that cannot be seen through the skin. Scientists tried to figure out what was inside an atom without being able to see inside the atom. In this lesson, you will learn the methods scientists used to "see" inside an atom.

Key Questions

What are three kinds of subatomic particles?

How can you describe the structure of the nuclear atom?

Vocabulary

- electron
- cathode ray
- proton
- neutron
- nucleus

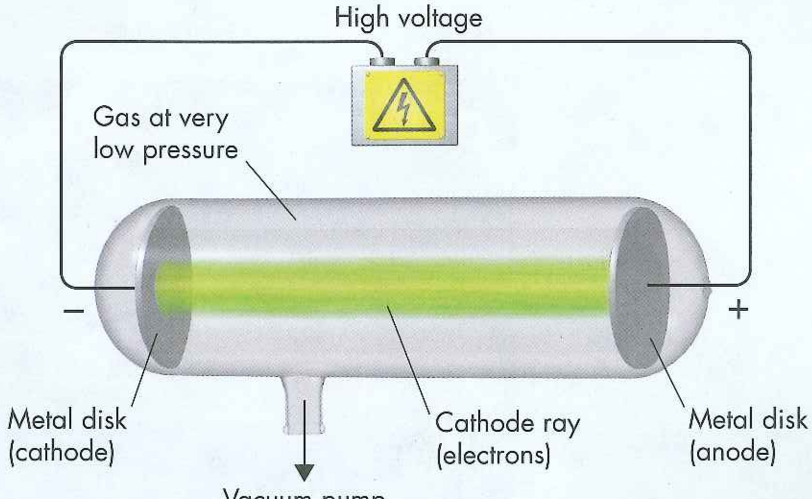
Subatomic Particles

What are three kinds of subatomic particles?

Much of Dalton's atomic theory is accepted today. One important change, however, is that atoms are now known to be divisible. They can be broken down into even smaller, more fundamental particles, called subatomic particles. **Three kinds of subatomic particles are electrons, protons, and neutrons.**

Electrons In 1897, the English physicist J. J. Thomson (1856–1940) discovered the electron. **Electrons** are negatively charged subatomic particles. Thomson performed experiments that involved passing electric current through gases at low pressure. He sealed the gases in glass tubes fitted at both ends with metal disks called electrodes. The electrodes were connected to a source of electricity, as shown in Figure 4.4. One electrode, the anode, became positively charged. The other electrode, the cathode, became negatively charged. The result was a glowing beam, or **cathode ray,** that traveled from the cathode to the anode.

Figure 4.4
Cathode-Ray Tube
In a cathode-ray tube, electrons travel as a ray from the cathode (−) to the anode (+). Televisions used to be made with a specialized type of cathode-ray tube.



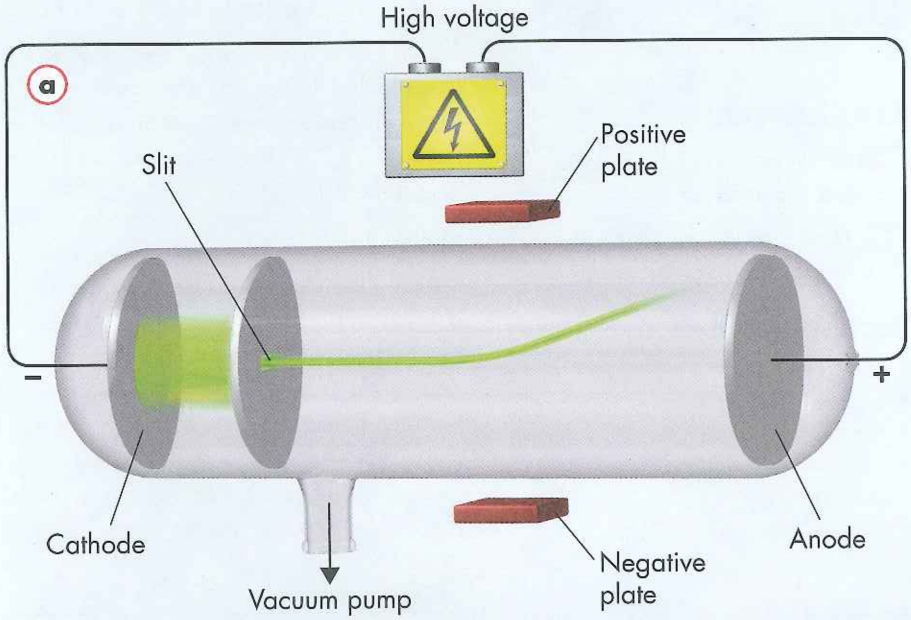
Thomson found that a cathode ray is deflected by electrically charged metal plates, as in Figure 4.5a. A positively charged plate attracts the cathode ray, while a negatively charged plate repels it. Thomson knew that opposite charges attract and like charges repel, so he hypothesized that a cathode ray is a stream of tiny negatively charged particles moving at high speed. Thomson called these particles corpuscles; later they were named electrons.

To test his hypothesis, Thomson set up an experiment to measure the ratio of an electron's charge to its mass. He found this ratio to be constant. Also, the charge-to-mass ratio of electrons did not depend on the kind of gas in the cathode-ray tube or the type of metal used for the electrodes. Thomson concluded that electrons are a component of the atoms of all elements.

The U.S. physicist Robert A. Millikan (1868–1953) carried out experiments to find the quantity of an electron's charge. In his oil-drop experiment, Millikan suspended negatively charged oil droplets between two charged plates. He then changed the voltage on the plates to see how this affected the droplets' rate of fall. From his data, he found that the charge on each oil droplet was a multiple of 1.60×10^{-19} coulomb, meaning this must be the charge of an electron. Using this charge value and Thomson's charge-to-mass ratio of an electron, Millikan calculated an electron's mass. Millikan's values for electron charge and mass are similar to those accepted today. An electron has one unit of negative charge, and its mass is 1/1840 the mass of a hydrogen atom.

Figure 4.5
Thomson's Experiment
a. Thomson found that cathode rays are attracted to metal plates that have a positive electrical charge. **b.** A cathode ray can also be deflected by a magnet.
Infer ***If a cathode ray is attracted to a positively charged plate, what can you infer about the charge of the particles that make up the cathode ray?***

KINETIC ART **See cathode-ray tubes animated online.**



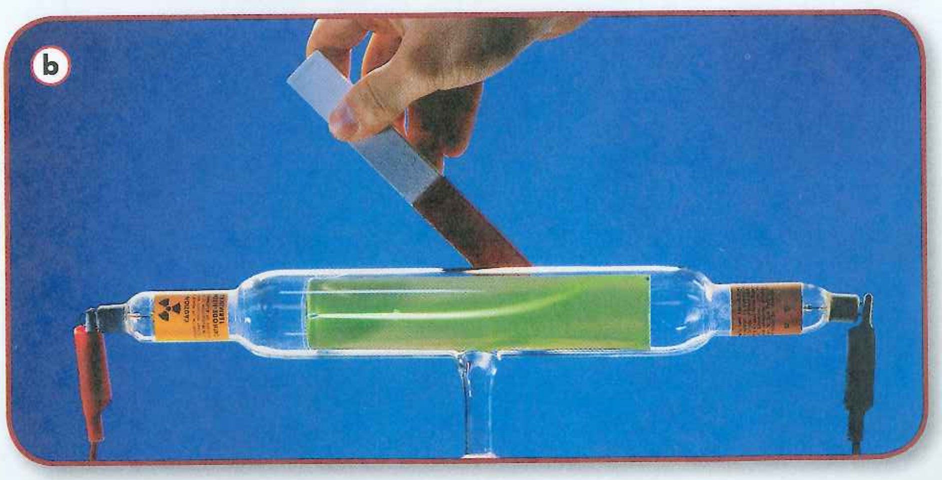


Table 4.1

Properties of Subatomic Particles				
Particle	Symbol	Relative charge	Relative mass (mass of proton = 1)	Actual mass (g)
Electron	e^-	1–	1/1840	9.11×10^{-28}
Proton	p^+	1+	1	1.67×10^{-24}
Neutron	n^0	0	1	1.67×10^{-24}

Protons and Neutrons If cathode rays are electrons given off by atoms, what remains of the atoms that have lost the electrons? For example, after a hydrogen atom (the lightest kind of atom) loses an electron, what is left? You can think through this problem using four simple ideas about matter and electric charges. First, atoms have no net electric charge; they are electrically neutral. (One important piece of evidence for electrical neutrality is that you do not receive an electric shock every time you touch something!) Second, electric charges are carried by particles of matter. Third, electric charges always exist in whole-number multiples of a single basic unit; that is, there are no fractions of charges. Fourth, when a given number of negatively charged particles combines with an equal number of positively charged particles, an electrically neutral particle is formed.

Considering all of this information, it follows that a particle with one unit of positive charge should remain when a typical hydrogen atom loses an electron. Evidence for such a positively charged particle was found in 1886, when Eugen Goldstein (1850–1930) observed a cathode-ray tube and found rays traveling in the direction opposite to that of the cathode rays. He called these rays canal rays and concluded that they were composed of positive particles. Such positively charged subatomic particles are called **protons.** Each proton has a mass about 1840 times that of an electron.

In 1932, the English physicist James Chadwick (1891–1974) confirmed the existence of yet another subatomic particle: the neutron. **Neutrons** are subatomic particles with no charge but with a mass nearly equal to that of a proton. Table 4.1 summarizes the properties of these subatomic particles. Although protons and neutrons are exceedingly small, theoretical physicists believe that they are composed of yet smaller subnuclear particles called *quarks*.

The Atomic Nucleus

How can you describe the structure of the nuclear atom?

When subatomic particles were discovered, scientists wondered how the particles were put together in an atom. This question was difficult to answer, given how tiny atoms are. Most scientists—including J. J. Thomson, discoverer of the electron—thought it likely that electrons were evenly distributed throughout an atom filled uniformly with positively charged material. In Thomson's atomic model, known as the "plum-pudding model," electrons were stuck into a lump of positive charge, similar to raisins stuck in dough. This model of the atom turned out to be short-lived, however, due to the work of a former student of Thomson, Ernest Rutherford (1871–1937), shown in Figure 4.6.

Figure 4.6 Ernest Rutherford
Born in New Zealand, Rutherford was awarded the Nobel Prize in Chemistry in 1908. His portrait appears on the New Zealand $100 bill.

CHEMISTRY & YOU

Q: *How did scientists "see" inside an atom to determine the structures that are inside an atom?*

Rutherford's Gold-Foil Experiment In 1911, Rutherford and his co-workers at the University of Manchester, England, wanted to test the existing plum-pudding model of atomic structure. So, they devised the gold-foil experiment. Their test used alpha particles, which are helium atoms that have lost their two electrons and have a double positive charge because of the two remaining protons. In the experiment, illustrated in Figure 4.7, a narrow beam of alpha particles was directed at a very thin sheet of gold foil. According to the prevailing theory, the alpha particles should have passed easily through the gold, with only a slight deflection due to the positive charge thought to be spread out in the gold atoms.

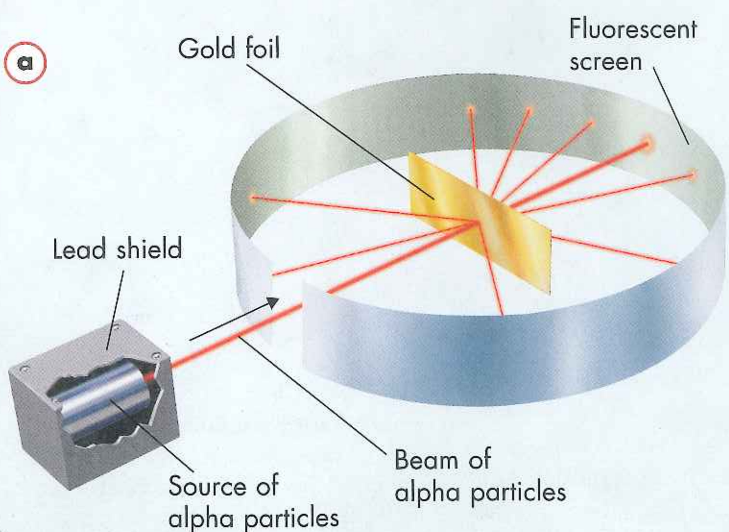
Rutherford's results were that most alpha particles went straight through the gold foil, or were slightly deflected. However, what was surprising is that a small fraction of the alpha particles bounced off the gold foil at very large angles. Some even bounced straight back toward the source. Rutherford later recollected, "This is almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you."

The Rutherford Atomic Model Based on his experimental results, Rutherford suggested a new theory of the atom. He proposed that the atom is mostly empty space, thus explaining the lack of deflection of most of the alpha particles. He concluded that all the positive charge and almost all the mass are concentrated in a small region that has enough positive charge to account for the great deflection of some of the alpha particles. He called this region the nucleus. The **nucleus** is the tiny central core of an atom and is composed of protons and neutrons.

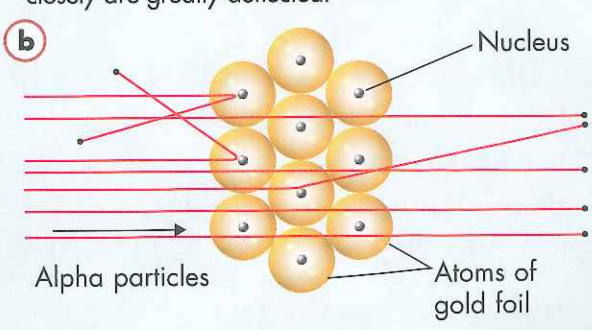
VIRTUAL LAB

Figure 4.7 Rutherford's Experiment
Rutherford's gold-foil experiment yielded evidence of the atomic nucleus.

a. Rutherford and his co-workers aimed a beam of alpha particles at a sheet of gold foil surrounded by a fluorescent screen. Most of the particles passed through the foil with no deflection at all. A few particles were greatly deflected.



b. Rutherford concluded that most of the alpha particles pass through the gold foil because the atom is mostly empty space. The mass and positive charge are concentrated in a small region of the atom. Rutherford called this region the nucleus. Particles that approach the nucleus closely are greatly deflected.



KINETIC ART *See* **Rutherford's gold-foil experiment** *animated online.*

The Rutherford atomic model is known as the nuclear atom. **In the nuclear atom, the protons and neutrons are located in the positively charged nucleus. The electrons are distributed around the nucleus and occupy almost all the volume of the atom.** According to this model, the nucleus is tiny and densely packed compared with the atom as a whole. If an atom were the size of a football stadium, the nucleus would be about the size of a marble.

Although it was an improvement over Thomson's model of the atom, Rutherford's model turned out to be incomplete. In Chapter 5, you will learn how the Rutherford atomic model had to be revised in order to explain the chemical properties of elements.

Quick Lab

Using Inference: The Black Box

Purpose To determine the shape of a fixed object inside a sealed box without opening the box

Materials

- **box containing a regularly shaped object fixed in place and a loose marble**

Procedure

1. Do not open the box.
2. Manipulate the box so that the marble moves around the fixed object.
3. Gather data (clues) that describe the movement of the marble.
4. Sketch a picture of the object in the box, showing its shape, size, and location within the box.
5. Repeat this activity with a different box containing a different object.

Analyze and Conclude

1. **Compare** Find a classmate who had a box with the same letter as yours, and compare your findings.
2. **Apply Concepts** Think about the experiments that have contributed to a better understanding of the atom. Which experiment does this activity remind you of?

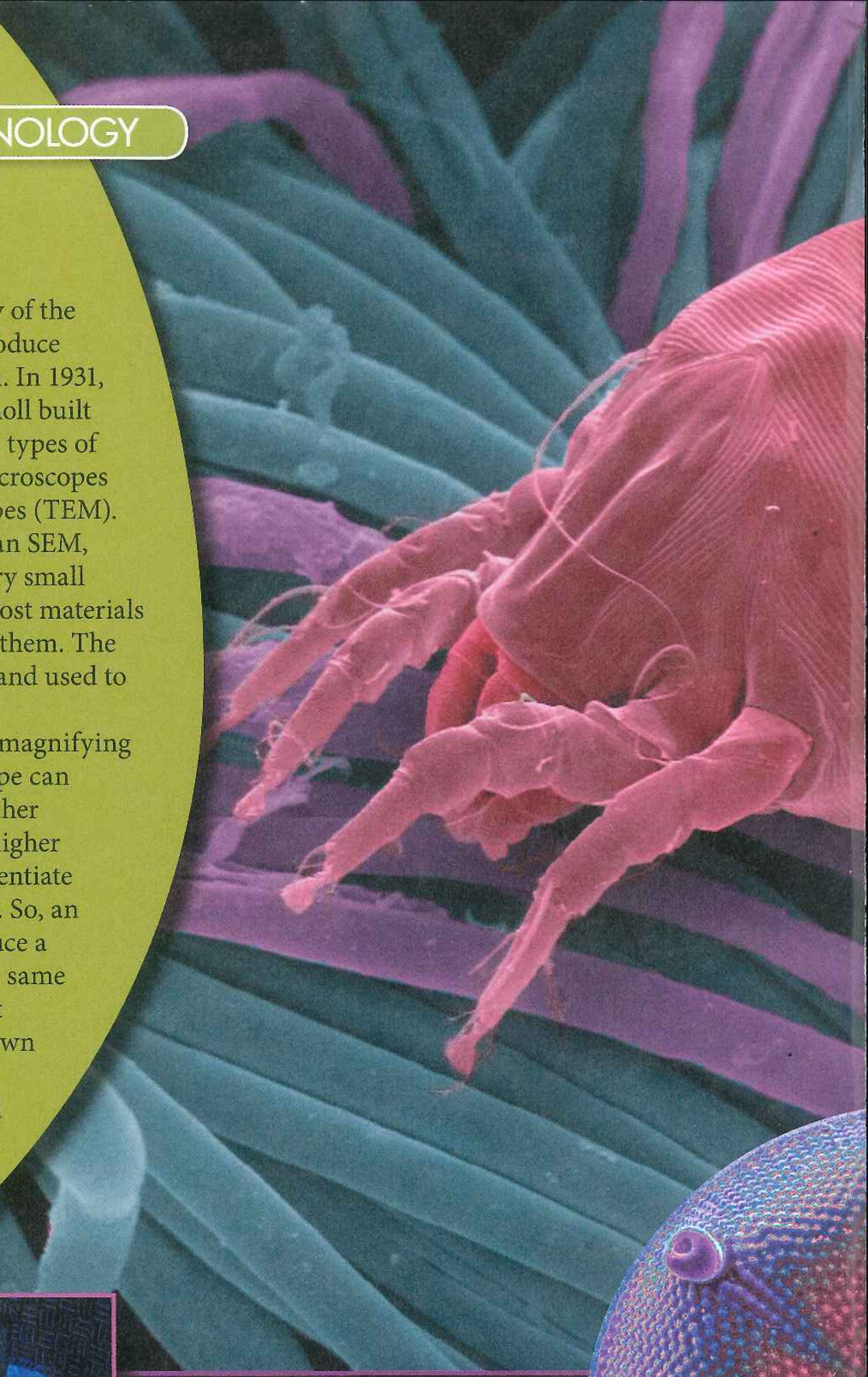
4.2 LessonCheck

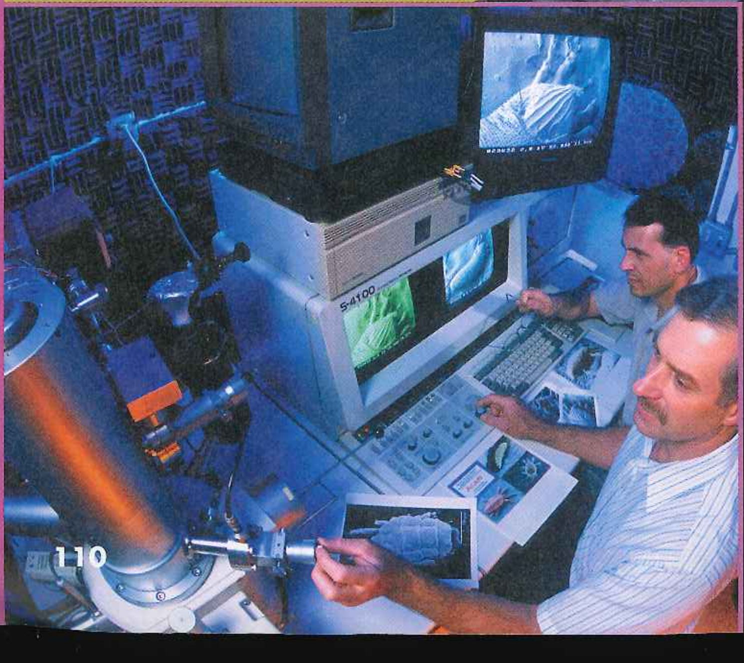
9. **Review** What are three types of subatomic particles?
10. **Explain** How does the Rutherford model describe the structure of atoms?
11. **Review** What are the charges and relative masses of the three main subatomic particles?
12. **Explain** Describe Thomson's and Millikan's contributions to atomic theory.
13. **Compare and Contrast** Compare Rutherford's expected outcome of the gold-foil experiment with the actual outcome.
14. **Analyze Data** What experimental evidence led Rutherford to conclude that an atom is mostly empty space?
15. **Compare and Contrast** How did Rutherford's model of the atom differ from Thomson's?

Electron Microscopy

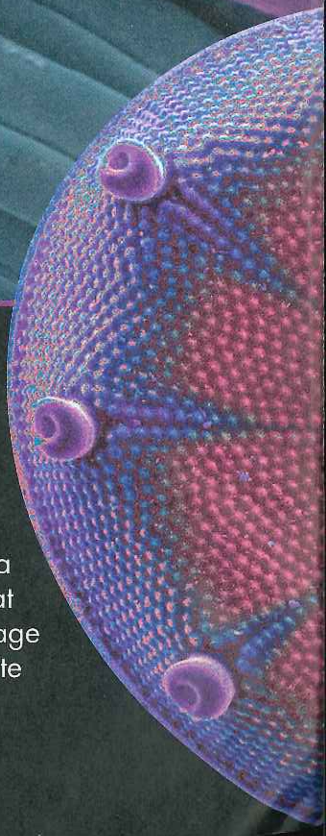
Within 30 years of J. J. Thomson's discovery of the electron, scientists were studying how to produce images of objects by using an electron beam. In 1931, German scientists Ernst Ruska and Max Knoll built the first electron microscope. There are two types of electron microscopes, scanning electron microscopes (SEM) and transmission electron microscopes (TEM). The images shown here are from SEMs. In an SEM, a beam of electrons is focused down to a very small diameter and scanned across the sample. Most materials eject electrons when the electron beam hits them. The location of the ejected electrons is detected and used to produce an image.

A typical light microscope is capable of magnifying an object 1000 times. An electron microscope can magnify an object over 100,000 times. Another advantage of electron microscopes is their higher resolution. Resolution is the ability to differentiate two objects that are very close to each other. So, an electron microscope has the ability to produce a clearer image than a light microscope at the same magnification. Electron microscopes do not produce color images. The color images shown here have false color that has been added to the images. Electron microscopes are useful in chemistry, but also in other fields, such as archeolology, pharmacology and quality assurance testing.

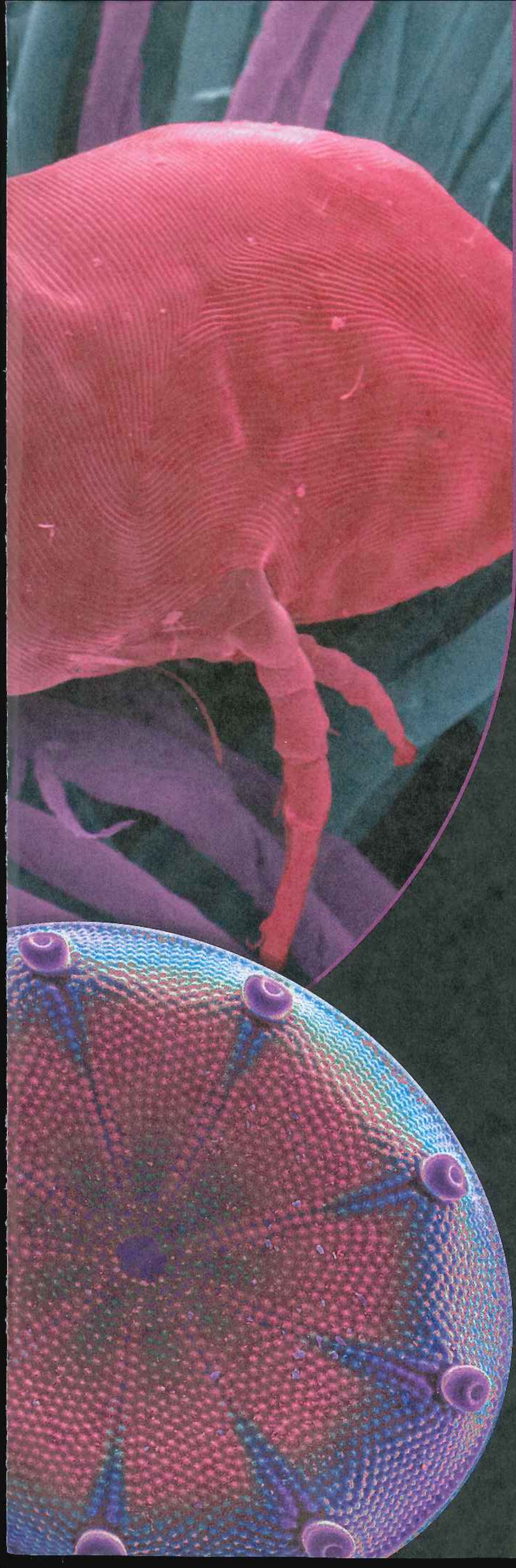


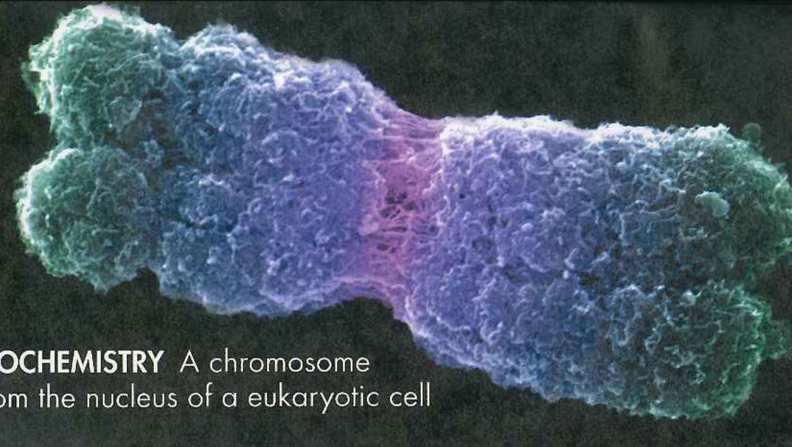


SEM This microscope is a scanning electron microscope. The image from the microscope is seen by using a computer screen.

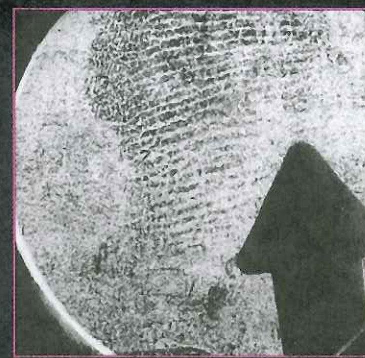


BIOLOGY This diatom is a single-celled organism that lives in the water. The image shown above is a dust mite on a piece of fabric.

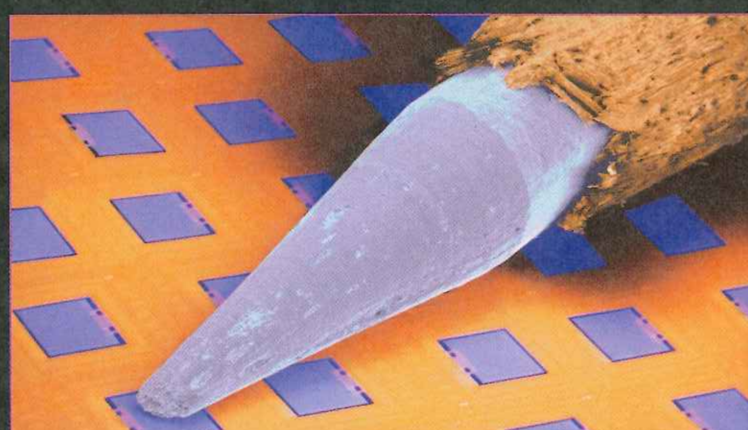




BIOCHEMISTRY A chromosome from the nucleus of a eukaryotic cell



FORENSICS The image on the left is from a light microscope, the image on the right from an SEM. The SEM image shows a clearer image of the fingerprint left on the page because oil from our fingers produces a different intensity of ejected electrons than paper or the ink.



MATERIALS SCIENCE A tip of a pencil is pointing to a tiny pressure sensor (mauve square) used in a car's air bag. When the sensors detect rapid deceleration, they trigger the inflation of the air bag.

Take It Further

1. Infer Why would a forensic investigator want to analyze gunshot residue using an electron microscope?

2. Compare Research the differences between an SEM and a TEM.

4.3 Distinguishing Among Atoms

CHEMISTRY & YOU

Q: *How can there be different varieties of atoms?* Some things exist in many different varieties. For example, dogs can differ in many ways, such as color, size, ear shape, and length of hair. Just as there are many types of dogs, atoms come in different varieties, too.

Key Questions

What makes one element different from another?

How do isotopes of an element differ?

How do you calculate the atomic mass of an element?

Vocabulary

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

Atomic Number and Mass Number

What makes one element different from another?

Atoms are composed of protons, neutrons, and electrons. Protons and neutrons make up the nucleus. Electrons surround the nucleus. How, then, are atoms of hydrogen, for example, different from atoms of oxygen?

Atomic Number Look at Table 4.2. Notice that a hydrogen atom has one proton, but an oxygen atom has eight protons. **Elements are different because they contain different numbers of protons.** An element's **atomic number** is the number of protons in the nucleus of an atom of that element. Since all hydrogen atoms have one proton, the atomic number of hydrogen is 1. All oxygen atoms have eight protons, so the atomic number of oxygen is 8. The atomic number identifies an element. For each element listed in Table 4.2, the number of protons equals the number of electrons. Remember that atoms are electrically neutral. Thus, the number of electrons (negatively charged particles) must equal the number of protons (positively charged particles).

Table 4.2

Atoms of the First Ten Elements

Name	Symbol	Atomic number	Protons	Neutrons*	Mass number	Electrons
Hydrogen	H	1	1	0	1	1
Helium	He	2	2	2	4	2
Lithium	Li	3	3	4	7	3
Beryllium	Be	4	4	5	9	4
Boron	B	5	5	6	11	5
Carbon	C	6	6	6	12	6
Nitrogen	N	7	7	7	14	7
Oxygen	O	8	8	8	16	8
Fluorine	F	9	9	10	19	9
Neon	Ne	10	10	10	20	10

* Number of neutrons in the most abundant isotope. Isotopes are introduced later in Lesson 4.3.

Sample Problem 4.1

Understanding Atomic Number

The element nitrogen (N) has an atomic number of 7. How many protons and electrons are in a neutral nitrogen atom?

1 Analyze **Identify the relevant concepts.** The atomic number gives the number of protons, which in a neutral atom equals the number of electrons.

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

Identify the atomic number. Then use the atomic number to find the number of protons and electrons.

The atomic number of nitrogen is 7. So, a neutral nitrogen atom has 7 protons and 7 electrons.

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

a. fluorine (atomic number = 9)
b. calcium (atomic number = 20)
c. aluminum (atomic number = 13)
d. potassium (atomic number = 19)

17. Complete the table.

Element	Atomic number	Protons	Electrons
S	16	a. ___	b. ___
V	c. ___	23	d. ___
e. ___	f. ___	g. ___	5

Mass Number Most of the mass of an atom is concentrated in its nucleus and depends on the number of protons and neutrons. The total number of protons and neutrons in an atom is called the **mass number.** For example, a helium atom has two protons and two neutrons, so its mass number is 4. A carbon atom has six protons and six neutrons, so its mass number is 12.

If you know the atomic number and mass number of an atom of any element, you can determine the atom's composition. The number of neutrons in an atom is the difference between the mass number and atomic number.

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}$$

Table 4.2 shows that a fluorine atom has an atomic number of 9 and a mass number of 19. Since the atomic number equals the number of protons, which equals the number of electrons, a fluorine atom has nine protons and nine electrons. The mass number of fluorine is equal to the number of protons plus the number of neutrons. So the fluorine atom has ten neutrons, which is the difference between the mass number and the atomic number ($19 - 9 = 10$).

The composition of any atom can be represented in shorthand notation using the atomic number and mass number, as in Figure 4.8. The chemical symbol for gold, Au, appears with two numbers written to its left. The atomic number is the subscript. The mass number is the superscript. You can also refer to atoms by using the mass number and the name of the element. For example, $^{197}_{79}Au$ may be written as gold-197.

Figure 4.8 Chemical Symbol
Au is the chemical symbol for gold.
Apply Concepts *How many electrons does a gold atom have?*

CHEM TUTOR

Sample Problem 4.2

Determining the Composition of an Atom

How many protons, electrons, and neutrons are in each atom?

a. $^{9}_{4}Be$ **b.** $^{20}_{10}Ne$ **c.** $^{23}_{11}Na$

KNOWNS	UNKNOWNS
• Beryllium (Be) atomic number = 4 mass number = 9 • Neon (Ne) atomic number = 10 mass number = 20 • Sodium (Na) atomic number = 11 mass number = 23	number of: protons = ? electrons = ? neutrons = ?

1 Analyze **List the knowns and the unknowns.** Use the definitions of atomic number and mass number to calculate the numbers of protons, electrons, and neutrons.

2 Calculate **Solve for the unknowns.**

Use the atomic number to find the number of protons.

atomic number = number of protons

a. 4 **b.** 10 **c.** 11

Use the atomic number to find the number of electrons.

atomic number = number of electrons

a. 4 **b.** 10 **c.** 11

Use the mass number and atomic number to find the number of neutrons.

number of neutrons = mass number − atomic number

a. number of neutrons = 9 − 4 = 5

b. number of neutrons = 20 − 10 = 10

c. number of neutrons = 23 − 11 = 12

3 Evaluate **Do the results make sense?** For each atom, the mass number equals the number of protons plus the number of neutrons. The results make sense.

18. How many neutrons are in each atom?

a. $^{80}_{35}Br$ **b.** $^{32}_{16}S$ **c.** $^{108}_{47}Ag$ **d.** $^{207}_{82}Pb$

19. Use Table 4.2 to express the composition of each atom below in shorthand form.

a. carbon-12 **c.** beryllium-9

b. boron-11 **d.** oxygen-16

CHEMISTRY & YOU

Q: *How are atoms of one element different from the atoms of another element? How are isotopes of the same element different?*

Isotopes

How do isotopes of an element differ?

Figure 4.9 shows that there are three different kinds of neon atoms. How do these atoms differ? All have the same number of protons (10) and electrons (10), but they each have different numbers of neutrons. **Isotopes** are atoms that have the same number of protons but different numbers of neutrons. **Because isotopes of an element have different numbers of neutrons, they also have different mass numbers.** Despite these differences, isotopes are chemically alike because they have identical numbers of protons and electrons, which are the subatomic particles responsible for chemical behavior. Remember the dogs at the beginning of the lesson. Their color or size doesn't change the fact that they are all dogs. Similarly, the number of neutrons in isotopes of an element doesn't change which element it is because the atomic number doesn't change.

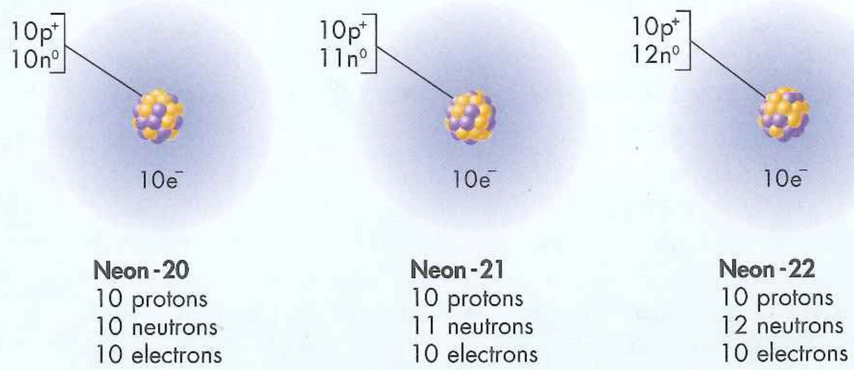


Figure 4.9 Isotopes Neon-20, neon-21, and neon-22 are three isotopes of neon. **Compare and Contrast** *How are these isotopes different? How are they similar?*

There are three known isotopes of hydrogen. Each isotope of hydrogen has one proton in its nucleus. The most common hydrogen isotope has no neutrons. It has a mass number of 1 and is called hydrogen-1 ($^{1}_{1}H$) or hydrogen. The second isotope has one neutron and a mass number of 2. It is called either hydrogen-2 ($^{2}_{1}H$) or deuterium. The third isotope has two neutrons and a mass number of 3. This isotope is called hydrogen-3 ($^{3}_{1}H$) or tritium.

SampleProblem 4.3

Writing Chemical Symbols for Isotopes

Diamonds are a naturally occurring form of elemental carbon. Two stable isotopes of carbon are carbon-12 and carbon-13. Write the symbol for each isotope using superscripts and subscripts to represent the mass number and the atomic number.

1 Analyze **Identify the relevant concepts.** Isotopes are atoms that have the same number of protons but different numbers of neutrons. The composition of an atom can be expressed by writing the chemical symbol, with the atomic number as a subscript and the mass number as a superscript.

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

Use Table 4.2 to identify the symbol and the atomic number for carbon.	The symbol for carbon is C. The atomic number of carbon is 6.
Look at the name of the isotope to find the mass number.	For carbon-12, the mass number is 12. For carbon-13, the mass number is 13.
Use the symbol, atomic number, and mass number to write the symbol of the isotope.	For carbon-12, the symbol is $^{12}_{6}C$. For carbon-13, the symbol is $^{13}_{6}C$.

20. Three isotopes of oxygen are oxygen-16, oxygen-17, and oxygen-18. Write the symbol for each, including the atomic number and mass number.

21. Three chromium isotopes are chromium-50, chromium-52, and chromium-53. How many neutrons are in each isotope, given that chromium has an atomic number of 24?

Atomic Mass

How do you calculate the atomic mass of an element?

A glance back at Table 4.1 on page 107 shows that the actual mass of a proton or a neutron is very small (1.67×10^{-24} g). The mass of an electron is 9.11×10^{-28} g, which is negligible in comparison. Given these values, the mass of even the largest atom is incredibly small. Since the 1920s, it has been possible to determine these tiny masses by using a mass spectrometer. With this instrument, the mass of a fluorine atom was found to be 3.155×10^{-23} g, and the mass of an arsenic atom was found to be 1.244×10^{-22} g. Such data about the actual masses of individual atoms can provide useful information, but in general, these values are inconveniently small and impractical to work with. Instead, it is more useful to compare the relative masses of atoms using a reference isotope as a standard. The reference isotope chosen is carbon-12. This isotope of carbon has been assigned a mass of exactly 12 atomic mass units. An **atomic mass unit (amu)** is defined as one twelfth of the mass of a carbon-12 atom. Using these units, a helium-4 atom has one third the mass of a carbon-12 atom. On the other hand, a nickel-60 atom has five times the mass of a carbon-12 atom.

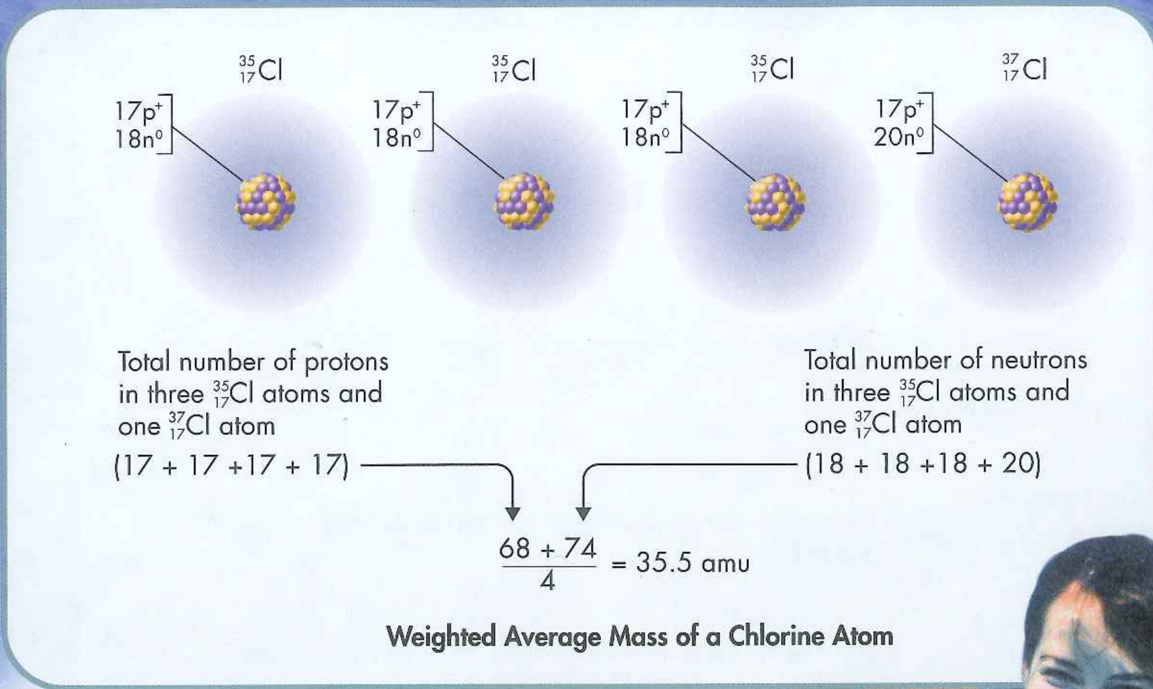
Interpret Data

Natural Percent Abundance of Stable Isotopes of Some Elements

Name	Symbol	Natural percent abundance	Mass (amu)	Atomic mass
Hydrogen	$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$	99.985 0.015 negligible	1.0078 2.0141 3.0160	1.0079
Helium	$^{3}_{2}He$ $^{4}_{2}He$	0.0001 99.9999	3.0160 4.0026	4.0026
Carbon	$^{12}_{6}C$ $^{13}_{6}C$	98.89 1.11	12.000 13.003	12.011
Nitrogen	$^{14}_{7}N$ $^{15}_{7}N$	99.63 0.37	14.003 15.000	14.007
Oxygen	$^{16}_{8}O$ $^{17}_{8}O$ $^{18}_{8}O$	99.759 0.037 0.204	15.995 16.995 17.999	15.999
Sulfur	$^{32}_{16}S$ $^{33}_{16}S$ $^{34}_{16}S$ $^{36}_{16}S$	95.002 0.76 4.22 0.014	31.972 32.971 33.967 35.967	32.06
Chlorine	$^{35}_{17}Cl$ $^{37}_{17}Cl$	75.77 24.23	34.969 36.966	35.453

Table 4.3 The atomic mass of an element is calculated using the percent abundance and mass of its isotopes.
a. Identify Which isotope of oxygen is the most abundant?
b. Describe How could you use the atomic mass of helium to determine which isotope of helium is most abundant?

Hint: The natural percent abundance of hydrogen-3 is "negligible" because the amount of naturally occurring hydrogen-3 is so small that it doesn't affect the atomic mass of hydrogen.



A carbon-12 atom has six protons and six neutrons in its nucleus, and its mass is set as 12 amu. The six protons and six neutrons account for nearly all of this mass. Therefore, the mass of a single proton or a single neutron is about one twelfth of 12 amu, or about 1 amu. Because the mass of any single atom depends mainly on the number of protons and neutrons in the nucleus of the atom, you might predict that the atomic mass of an element should be a whole number. However, that is not usually the case.

In nature, most elements occur as a mixture of two or more isotopes. Each isotope of an element has a fixed mass and a natural percent abundance. Consider the three isotopes of hydrogen discussed earlier in this section. According to Table 4.3, almost all naturally occurring hydrogen (99.985 percent) is hydrogen-1. The other two isotopes are present in trace amounts. Notice that the atomic mass of hydrogen listed in Table 4.3 (1.0079 amu) is very close to the mass of hydrogen-1 (1.0078 amu). The slight difference takes into account the larger masses, but much smaller amounts, of the other two isotopes of hydrogen.

Now consider the two stable isotopes of chlorine listed in Table 4.3: chlorine-35 and chlorine-37. If you calculate the arithmetic mean of these two masses ((34.969 amu + 36.966 amu)/2), you get an average atomic mass of 35.968 amu. However, this value is higher than the actual value of 35.453. To explain this difference, you need to know the natural percent abundance of the isotopes of chlorine. Chlorine-35 accounts for 75 percent of the naturally occurring chlorine atoms; chlorine-37 accounts for only 25 percent. See Figure 4.10. The **atomic mass** of an element is a weighted average mass of the atoms in a naturally occurring sample of the element. A weighted average mass reflects both the mass and the relative abundance of the isotopes as they occur in nature.

Figure 4.10 Isotopes of Chlorine
Chlorine is a reactive element used to disinfect swimming pools. Chlorine occurs as two isotopes: chlorine-35 and chlorine-37. Because there is more chlorine-35 than chlorine-37 in nature, the atomic mass of chlorine, 35.453 amu, is closer to 35 than to 37.
Evaluate ***How does a weighted average differ from an arithmetic mean?***

Sample Problem 4.4

Understanding Relative Abundance of Isotopes

The atomic mass of copper is 63.546 amu. Which of copper's two isotopes is more abundant: copper-63 or copper-65?

1 Analyze **Identify the relevant concepts.** The atomic mass of an element is the weighted average mass of the atoms in a naturally occurring sample of the element.

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

Compare the atomic mass to the mass of each isotope.

The atomic mass of 63.546 amu is closer to 63 than it is to 65.

Determine the most abundant isotope based on which isotope's mass is closest to the atomic mass.

Because the atomic mass is a weighted average of the isotopes, copper-63 must be more abundant than copper-65.

22. Boron has two isotopes: boron-10 and boron-11. Which is more abundant, given that the atomic mass of boron is 10.81 amu?

23. There are three isotopes of silicon; they have mass numbers of 28, 29, and 30. The atomic mass of silicon is 28.086 amu. Comment on the relative abundance of these three isotopes.

Now that you know that the atomic mass of an element is a weighted average of the masses of its isotopes, you can determine atomic mass based on relative abundance. To do this, you must know three things: the number of stable isotopes of the element, the mass of each isotope, and the natural percent abundance of each isotope. **To calculate the atomic mass of an element, multiply the mass of each isotope by its natural abundance, expressed as a decimal, and then add the products.** The resulting sum is the weighted average mass of the atoms of the element as they occur in nature. You can calculate the atomic masses listed in Table 4.3 based on the given masses and natural abundances of the isotopes for each element.

For example, carbon has two stable isotopes: carbon-12, which has a natural abundance of 98.89 percent, and carbon-13, which has a natural abundance of 1.11 percent. The mass of carbon-12 is 12.000 amu; the mass of carbon-13 is 13.003 amu. The atomic mass of carbon is calculated as follows:

$$\begin{aligned}\text{Atomic mass of carbon} &= (12.000\text{ amu} \times 0.9889) + (13.003\text{ amu} \times 0.0111)\\ &= (11.867\text{ amu}) + (0.144\text{ amu})\\ &= 12.011\text{ amu}\end{aligned}$$

Sample Problem 4.5

Calculating Atomic Mass

Element X has two naturally occurring isotopes. The isotope with a mass of 10.012 amu (^{10}X) has a relative abundance of 19.91 percent. The isotope with a mass of 11.009 amu (^{11}X) has a relative abundance of 80.09 percent. Calculate the atomic mass of element X.

To find all the knowns, change the percent abundance to decimals. A percent is a shorthand way of expressing a fraction whose denominator is 100. 19.91% is equivalent to 19.91/100 or 0.1991.

1 Analyze **List the knowns and the unknown.** The mass each isotope contributes to the element's atomic mass can be calculated by multiplying the isotope's mass by its relative abundance. The atomic mass of the element is the sum of these products.

KNOWNS

- isotope ^{10}X:
  mass = 10.012 amu
  relative abundance = 19.91% = 0.1991
- isotope ^{11}X:
  mass = 11.009 amu
  relative abundance = 80.09% = 0.8009

UNKNOWN

atomic mass of X = ?

2 Calculate **Solve for the unknown.**

Use the atomic mass and the decimal form of the percent abundance to the find the mass contributed by each isotope.

for ^{10}X: $10.012 \text{ amu} \times 0.1991 = 1.993 \text{ amu}$
for ^{11}X: $11.009 \text{ amu} \times 0.8009 = 8.817 \text{ amu}$

Add the atomic mass contributions for all the isotopes.

For element X, atomic mass $= 1.993 \text{ amu} + 8.817 \text{ amu}$
$= 10.810 \text{ amu}$

3 Evaluate **Does the result make sense?** The calculated value is closer to the mass of the more abundant isotope, as would be expected.

24. The element copper has naturally occurring isotopes with mass numbers of 63 and 65. The relative abundance and atomic masses are 69.2% for mass = 62.93 amu, and 30.8% for mass = 64.93 amu. Calculate the atomic mass of copper.

25. Calculate the atomic mass of bromine. The two isotopes of bromine have atomic masses and relative abundance of 78.92 amu (50.69%) and 80.92 amu (49.31%).

4.3 LessonCheck

26. **Explain** What distinguishes the atoms of one element from the atoms of another?

27. **Compare and Contrast** How do the isotopes of a given element differ from one another?

28. **Explain** How is atomic mass calculated?

29. **Identify** What equation tells you how to calculate the number of neutrons in an atom?

30. **Compare** How is atomic number different from mass number?

31. **Use Models** What does the number represent in the isotope platinum-194?

32. **Explain** The atomic masses of elements are generally not whole numbers. Explain why.

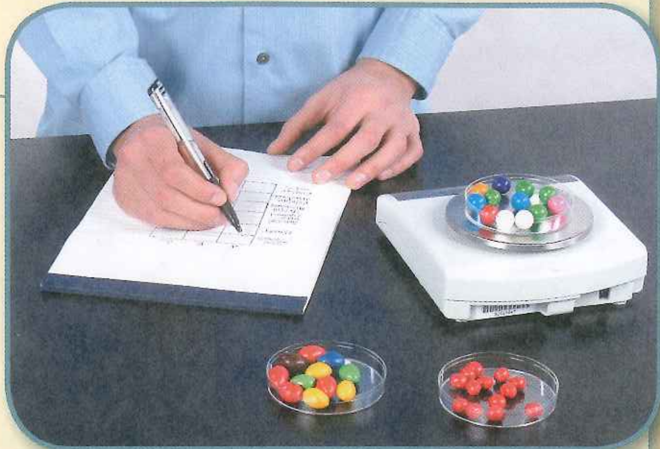
33. **Identify** Which of argon's three isotopes is most abundant: argon-36, argon-38, or argon-40? (*Hint:* the atomic mass of argon is 39.948 amu.)

34. **Calculate** List the number of protons, neutrons, and electrons in each pair of isotopes.

a. $^{6}_{3}Li$, $^{7}_{3}Li$ **b.** $^{42}_{20}Ca$, $^{44}_{10}Ca$ **c.** $^{78}_{34}Se$, $^{80}_{34}Se$

Small-Scale Lab

The Atomic Mass of "Candium"



Purpose

To analyze the isotopes of "candium" and to calculate its atomic mass

Materials

- sample of candium
- balance

Procedure

Obtain a sample of "candium" that contains three different brands of round, coated candy. Treat each brand of candy as an isotope of candium. Separate the three isotopes into groups labeled A, B, and C, and measure the mass of each isotope. Count the number of atoms in each sample. Make a table similar to the one below to record your measured and calculated data.

	A	B	C	Totals
Total mass (grams)				
Number				
Average mass (grams)				
Relative abundance				
Percent abundance				
Relative mass				

Analyze

Using the experimental data, record the answers to the following questions below your data table.

1. Calculate Calculate the average mass of each isotope by dividing its total mass by the number of particles of that isotope.

2. Calculate Calculate the relative abundance of each isotope by dividing its number of particles by the total number of particles.

3. Calculate Calculate the percent abundance of each isotope by multiplying the relative abundance from Step 2 by 100.

4. Calculate Calculate the relative mass of each isotope by multiplying its relative abundance from Step 2 by its average mass.

5. Calculate Calculate the weighted average mass of all candium particles by adding the relative masses. This weighted average mass is the atomic mass of candium.

6. Explain What is the difference between percent abundance and relative abundance? What is the result when you total the individual relative abundances? The individual percent abundances?

7. Identify The percent abundance of each kind of candy tells you how many of each kind of candy there are in every 100 particles. What does relative abundance tell you?

8. Analyze Data Compare the total values for rows 3 and 6 in the table. Explain why the totals differ and why the value in row 6 best represents atomic mass.

9. Analyze Data Explain any differences between the atomic mass of your candium sample and that of your neighbor. Explain why the difference would be smaller if larger samples were used.

You're the Chemist

The following small-scale activity allows you to develop your own procedures and analyze the results.

1. Analyze Data Determine the atomic mass of a second sample of candium. How does it compare with the first? Suggest reasons for any differences between the samples.

4 Study Guide

BIGIDEA ELECTRONS AND THE STRUCTURE OF ATOMS

Atoms are the smallest particles of an element that still have the chemical properties of that element. Atoms have positively charged protons and neutral neutrons inside a nucleus, and negatively charged electrons outside the nucleus. Atoms of the same element have the same number of protons, which is equal to an atom's atomic number. But atoms of the same element can have different numbers of neutrons. Atoms of the same element with different numbers of neutrons are isotopes.

4.1 Defining the Atom

Democritus reasoned that atoms were indivisible and indestructible. By using experimental methods, Dalton transformed Democritus's ideas on atoms into a scientific theory.

Scientists can observe individual atoms by using instruments such as scanning electron microscopes.

- atom (102)
- Dalton's atomic theory (103)

4.2 Structure of the Nuclear Atom

Three kinds of subatomic particles are electrons, protons, and neutrons.

In the nuclear atom, the protons and neutrons are located in the nucleus. The electrons are distributed around the nucleus and occupy almost all the volume of the atom.

- electron (105)
- cathode ray (105)
- proton (107)
- neutron (107)
- nucleus (108)

4.3 Distinguishing Among Atoms

Elements are different because they contain different numbers of protons.

Because isotopes of an element have different numbers of neutrons, they also have different mass numbers.

To calculate the atomic mass of an element, multiply the mass of each isotope by its natural abundance, expressed as a decimal, and then add the products.

- atomic number (112)
- mass number (113)
- isotope (114)
- atomic mass unit (amu) (116)
- atomic mass (117)

Key Equation

$$\text{number of neutrons} = \text{mass number} - \text{atomic number}$$

4 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

4.1 Defining the Atom

35. What is an atom?

36. What were the limitations of Democritus's ideas about atoms?

37. With which of these statements would John Dalton have agreed in the early 1800s? For each, explain why or why not.

a. Atoms are the smallest particles of matter.
b. The mass of an iron atom is different from the mass of a copper atom.
c. Every atom of silver is identical to every other atom of silver.
d. A compound is composed of atoms of two or more different elements.

38. Use Dalton's atomic theory to describe how atoms interact during a chemical reaction.

4.2 Structure of the Nuclear Atom

39. What experimental evidence did Thomson have for each statement?

a. Electrons have a negative charge.
b. Atoms of all elements contain electrons.

★**40.** Would you expect two electrons to attract or repel each other?

41. How do the charge and mass of a neutron compare to the charge and mass of a proton?

42. Why does it make sense that if an atom loses electrons, it is left with a positive charge?

43. Describe the location of the electrons in Thomson's "plum-pudding" model of the atom.

★**44.** How did the results of Rutherford's gold-foil experiment differ from his expectations?

45. What is the charge, positive or negative, of the nucleus of every atom?

46. In the Rutherford atomic model, which subatomic particles are located in the nucleus?

4.3 Distinguishing Among Atoms

47. Why is an atom electrically neutral?

48. What does the atomic number of each atom represent?

49. How many protons are in the nuclei of the following atoms?

a. phosphorus (P)
b. molybdenum (Mo)
c. aluminum (Al)
d. cadmium (Cd)
e. chromium (Cr)
f. lead (Pb)

50. What is the difference between the mass number and the atomic number of an atom?

★**51.** Complete the following table.

Atomic number	Mass number	Number of protons	Number of neutrons
9	a. ____	b. ____	10
c. ____	d. ____	14	15
e. ____	47	f. ____	25
g. ____	55	25	h. ____

52. Name two ways that isotopes of an element differ.

★**53.** Lithium has two isotopes, lithium-6 (atomic mass = 6.015, relative abundance = 7.5%) and lithium-7 (atomic mass = 7.016, relative abundance = 92.5%). Calculate the atomic mass of lithium.

Understand Concepts

★**54.** How can there be more than 1000 different atoms when there are only about 100 different elements?

55. What data must you know about the isotopes of an element to calculate the atomic mass of the element?

56. How is an average mass different from a weighted average mass?

57. What is the atomic mass of an element?

58. Characterize the size of an atom.

59. Compare the size and density of an atom with its nucleus.

★**60.** You are standing on the top of a boron-11 nucleus. Describe the numbers and kinds of subatomic particles you see looking down into the nucleus, and those you see looking out from the nucleus.

61. What parts of Dalton's atomic theory no longer agree with the current picture of the atom?

62. Millikan measured the quantity of charge carried by an electron. How did he then calculate the mass of an electron?

63. How is the atomic mass of an element calculated from isotope data?

***64.** The four isotopes of lead are shown below, each with its percent by mass abundance and the composition of its nucleus. Using these data, calculate the approximate atomic mass of lead.

65. Dalton's atomic theory was not correct in every detail. Should this be taken as a criticism of Dalton as a scientist? Explain.

***66.** The following table shows some of the data collected by Rutherford and his colleagues during their gold-foil experiment.

Angle of deflection (degrees)	Number of deflections
5	8,289,000
10	502,570
15	120,570
30	7800
45	1435
60	477
75	211
>105	198

a. What percentage of the alpha particle deflections were 5° or less?

b. What percentage of the deflections were 15° or less?

c. What percentage of the deflections were 60° or greater?

***67.** Using the data for nitrogen listed in Table 4.3, calculate the weighted average atomic mass of nitrogen. Show your work.

68. What characteristics of cathode rays led Thomson to conclude that the rays consisted of negatively charged particles?

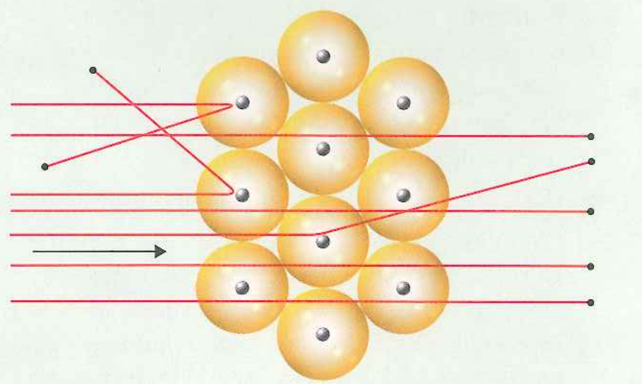
69. If you know the atomic number and mass number of an atom of an element, how can you determine the number of protons, neutrons, and electrons in that atom?

70. What makes isotopes of the same element chemically alike?

71. If isotopes are chemically alike, but physically different, propose which subatomic particles are responsible for determining an element's chemical reactivity.

Think Critically

72. Interpret Diagrams The diagram below shows gold atoms being bombarded with fast-moving alpha particles.



a. The large yellow spheres represent gold atoms. What do the small gray spheres represent?

b. List at least two characteristics of the small gray spheres.

c. Which subatomic particle cannot be found in the area represented by the gray spheres?

73. Evaluate and Revise How could you modify Rutherford's experimental procedure to determine the relative sizes of different nuclei?

*74. **Explain** Rutherford's atomic theory proposed a dense nucleus surrounded by very small electrons. This structure implies that atoms are composed mainly of empty space. If all matter is mainly empty space, why is it impossible to walk through walls or pass your hand through your desk?

75. **Explain** This chapter illustrates the scientific method in action. What happens when new experimental results cannot be explained by the existing theory?

76. **Apply Concepts** The law of conservation of mass was introduced in Chapter 2. Use Dalton's atomic theory to explain this law.

77. **Infer** Diamond and graphite are both composed of carbon atoms. The density of diamond is 3.52 g/cm^3. The density of graphite is 2.25 g/cm^3. In 1955, scientists successfully made diamond from graphite. Using the relative densities, consider what happens at the atomic level when this change occurs. Then suggest how this synthesis may have been accomplished.

*78. **Calculate** Lithium has two naturally occurring isotopes. Lithium-6 has an atomic mass of 6.015 amu; lithium-7 has an atomic mass of 7.016 amu. The atomic mass of lithium is 6.941 amu. What is the percentage of naturally occurring lithium-7?

*79. **Calculate** When the masses of the particles that make up an atom are added together, the sum is always larger than the actual mass of the atom. The missing mass, called the mass defect, represents the matter converted into energy when the nucleus was formed from its component protons and neutrons. Calculate the mass defect of a chlorine-35 atom by using the data in Table 4.1. The actual mass of a chlorine-35 atom is 5.81×10^{-23} g.

Write About Science

80. **Communicate** Explain how Rutherford's gold-foil experiment yielded new evidence about atomic structure. *Hint:* First, describe the setup of the experiment. Then, explain how Rutherford interpreted his experimental data.

81. **Connect to the BIGIDEA** Choose two atoms from Table 4.2. Compare and contrast the structure of the two atoms.

CHEMYSTERY

Artifacts or Artifakes?

There are currently no crystal skulls that have been proven to be from ancient civilizations. The Linnean Society of London is a research institute specializing in taxonomy and natural history. They have used electron microscopy to view the surface of crystal skulls, including a crystal skull that is part of the Smithsonian collection. Images from the scanning electron microscope reveal circular patterns on the surface of the skulls. These patterns indicate the skulls were likely carved using a modern carving device with a rotary wheel. Ancient civilizations would not have had such devices. Therefore, all the crystal skulls that are known today appear to be "artifakes," not artifacts.

82. **Infer** Why would electron microscopes be able to provide more information about an object than a light microscope?

83. **Connect to the BIGIDEA** How has knowledge of atomic structure aided in the development of the electron microscope?

Cumulative Review

84. How does a scientific law differ from a scientific theory?

85. Classify each as an element, a compound, or a mixture.
  a. sulfur
  b. salad oil
  c. newspaper
  d. orange

*86. Oxygen and hydrogen react explosively to form water. In one reaction, 6 g of hydrogen combines with oxygen to form 54 g of water. How much oxygen was used?

87. An aquarium measures 54.0 cm × 31.10 m × 80.0 cm. How many cubic centimeters of water will this aquarium hold?

*88. What is the mass of 4.42 cm^3 of platinum? The density of platinum is 22.5 g/cm^3.

If You Have Trouble With . . .

Question	84	85	86	87	88
See Chapter	1	2	2	3	3

Standardized Test Prep

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

1. The smallest particle of an element that retains its identity in a chemical reaction is a
   (A) proton.
   (B) neutron.
   (C) atom.
   (D) compound.

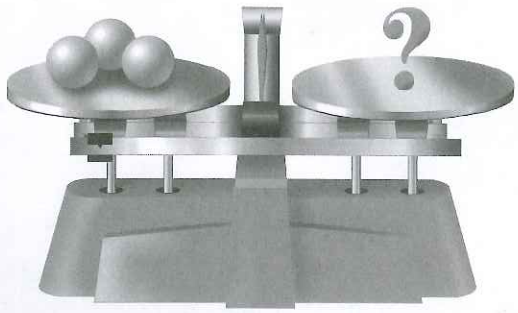
2. Which of these descriptions is *incorrect*?
   (A) proton: positive charge, in nucleus, mass of $\approx$1 amu
   (B) electron: negative charge, mass of $\approx$0 amu, in nucleus
   (C) neutron: mass of $\approx$1 amu, no charge

3. Thallium has two isotopes, thallium-203 and thallium-205. Thallium's atomic number is 81, and its atomic mass is 204.38 amu. Which statement about the thallium isotopes is true?
   (A) There is more thallium-203 in nature.
   (B) Atoms of both isotopes have 81 protons.
   (C) Thallium-205 atoms have fewer neutrons.
   (D) The most common atom of thallium has a mass of 204.38 amu.

4. Which atom is composed of 16 protons, 16 electrons, and 16 neutrons?
   (A) $^{48}_{16}S$ (C) $^{32}_{16}S$
   (B) $^{16}_{32}Ge$ (D) $^{16}_{32}S$

Use the art to answer Question 5.

5. How many nitrogen-14 atoms (^{14}N) would you need to place on the right pan to balance the three calcium-42 atoms (^{42}Ca) on the left pan of the "atomic balance" below? Describe the method you used to determine your answer, including any calculations.



Tips for Success

Connectors Sometimes two phrases in a true/false question are connected by a word such as *because* or *therefore*. These words imply a relationship between one part of the sentence and another. Statements that include such words can be false even if both parts of the statement are true by themselves.

For each question below, there are two statements. Decide whether each statement is true or false. Then decide whether Statement II is a correct explanation for Statement I.

Statement I		Statement II
6. Every aluminum-27 atom has 27 protons and 27 electrons.	BECAUSE	The mass number of aluminum-27 is 27.
7. Isotopes of an element have different atomic masses.	BECAUSE	The nuclei of an element's isotopes contain different numbers of protons.
8. An electron is repelled by a negatively charged particle.	BECAUSE	An electron has a negative charge.
9. In an atom, the number of neutrons is generally equal to or greater than the number of protons.	BECAUSE	The mass number is generally equal to or greater than the atomic number.

If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9
See Lesson	4.1	4.2	4.3	4.3	4.3	4.3	4.3	4.2	4.2