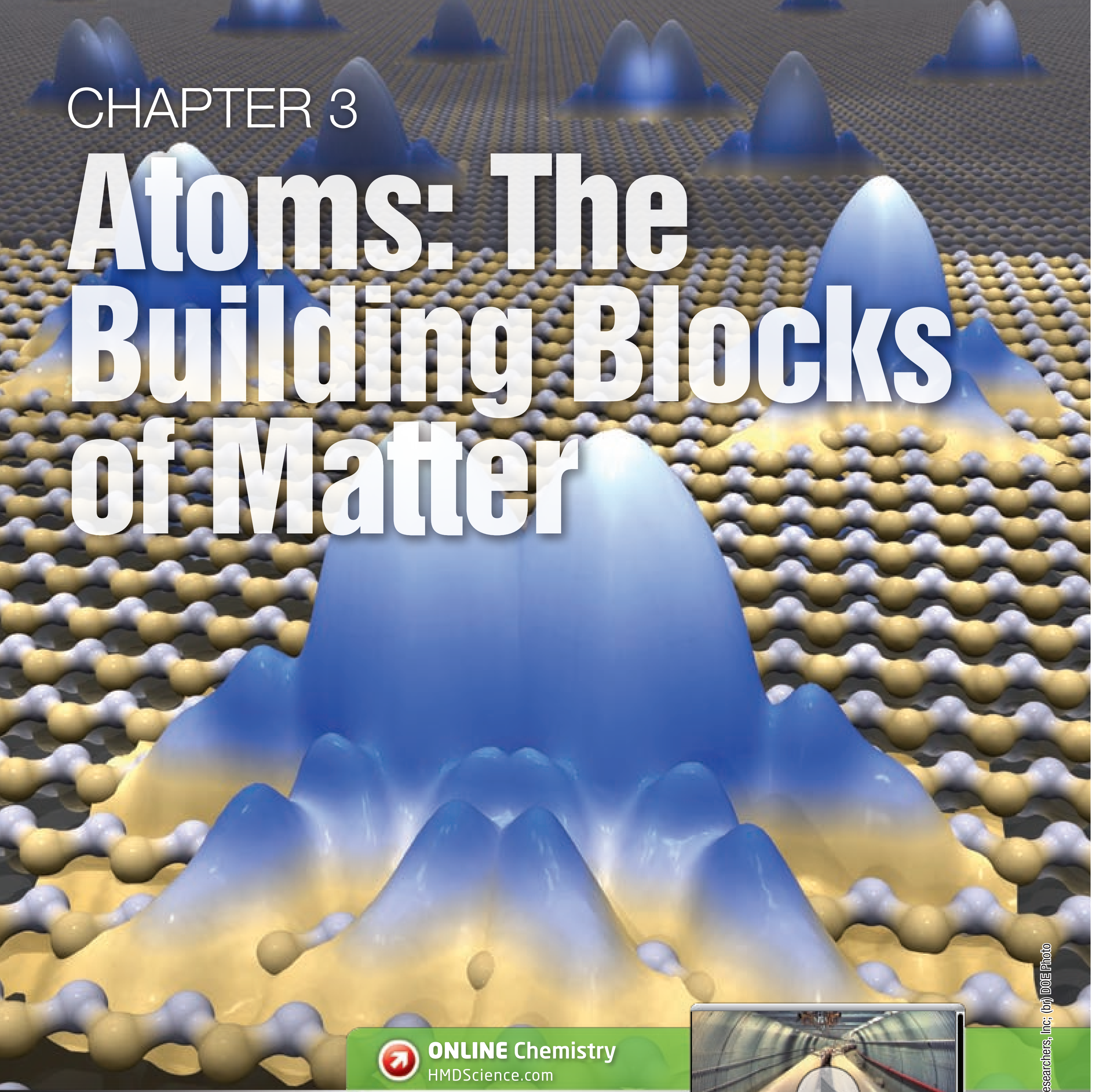
CHAPTER 3

Atoms: The Building Blocks of Matter



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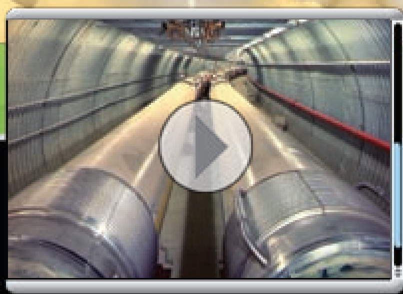

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Atoms

The Atom: From Philosophical Idea to Scientific Theory

SECTION 1

Main Ideas

- Three basic laws describe how matter behaves in chemical reactions.
- Compounds contain atoms in whole-number ratios.
- Atoms can be subdivided into smaller particles.

Key Terms

law of conservation of mass
law of definite proportions
law of multiple proportions

VIRGINIA STANDARDS

CH.2.i The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of: historical and quantum models.
CH.1.EKS-16; CH.1.EKS-26; CH.2.EKS-16

When you crush a lump of sugar, you can see that it is made up of many smaller particles of sugar. You may grind these particles into a very fine powder, but each tiny piece is still sugar. Now suppose you dissolve the sugar in water. The tiny particles seem to disappear completely. Even if you look at the sugar-water solution through a powerful microscope, you cannot see any sugar particles. Yet if you were to taste the solution, you'd know that the sugar is still there. Observations like these led early philosophers to ponder the fundamental nature of matter. Is it continuous and infinitely divisible, or is it divisible only until a basic, invisible particle that cannot be divided further is reached?

The particle theory of matter was supported as early as 400 BCE by certain Greek thinkers, such as Democritus. He called nature's basic particle an *atom*, based on the Greek word meaning "indivisible." Aristotle was part of the generation that succeeded Democritus. His ideas had a lasting impact on Western civilization, and he did not believe in atoms. He thought that all matter was continuous, and his opinion was accepted for nearly 2000 years. Neither the view of Aristotle nor that of Democritus was supported by experimental evidence, so each remained under speculation until the eighteenth century. Then scientists began to gather evidence favoring the atomic theory of matter.

MAIN IDEA

Three basic laws describe how matter behaves in chemical reactions.

Virtually all chemists in the late 1700s accepted the modern definition of an element as a substance that cannot be further broken down by ordinary chemical means. They also assumed that these elements combined to form compounds that have different physical and chemical properties than those of the elements that make them. What troubled them, however, was the understanding of just exactly how the different substances could combine with one another to form new ones, what we know as *chemical reactions*. Most historians date the foundation of modern chemistry to this time when scientists finally began to ascribe rules to how matter interacts.

FIGURE 1.1

Table Salt Crystals Each of the salt crystals shown here contains exactly 39.34% sodium and 60.66% chlorine by mass.

In the 1790s, the study of matter was revolutionized by a new emphasis on the quantitative analysis of chemical reactions. Aided by improved balances, investigators began to accurately measure the masses of the elements and compounds they were studying. This led to the discovery of several basic laws. **One of these laws was the law of conservation of mass, which states that mass is neither created nor destroyed during ordinary chemical reactions or physical changes.** This discovery was soon followed by the assertion that, regardless of where or how a pure chemical compound is prepared, it is composed of a fixed proportion of elements. For example, sodium chloride, also known as ordinary table salt, as shown in **Figure 1.1,** *always* consists of 39.34% by mass of the element sodium, Na, and 60.66% by mass of the element chlorine, Cl. **The fact that a chemical compound contains the same elements in exactly the same proportions by mass regardless of the size of the sample or source of the compound is known as the law of definite proportions.**

It was also known that two elements sometimes combine to form more than one compound. For example, the elements carbon and oxygen form two compounds, carbon dioxide and carbon monoxide. Consider samples of each of these compounds, each containing 1.00 g of carbon. In carbon dioxide, 2.66 g of oxygen combine with 1.00 g of carbon. In carbon monoxide, 1.33 g of oxygen combine with 1.00 g of carbon. The ratio of the masses of oxygen in these two compounds is 2.66 to 1.33, or 2 to 1. **This illustrates the law of multiple proportions: If two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers.**

MAIN IDEA

Compounds contain atoms in whole-number ratios.

In 1808, an English schoolteacher named John Dalton proposed an explanation that encompassed all these laws. He reasoned that elements were composed of atoms and that only whole numbers of atoms can combine to form compounds. His theory can be summed up by the following statements.

1. All matter is composed of extremely small particles called atoms.
2. Atoms of an element are identical in size, mass, and other properties; atoms of different elements differ in size, mass, and other properties.
3. Atoms cannot be subdivided, created, or destroyed.
4. Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
5. In chemical reactions, atoms are combined, separated, or rearranged.

Dalton's atomic theory explains the law of conservation of mass through the concept that chemical reactions involve merely the combination, separation, or rearrangement of atoms and that during reactions atoms are not subdivided, created, or destroyed. **Figure 1.2,** on the next page, illustrates this idea for the formation of carbon monoxide from carbon and oxygen.

FIGURE 1.2

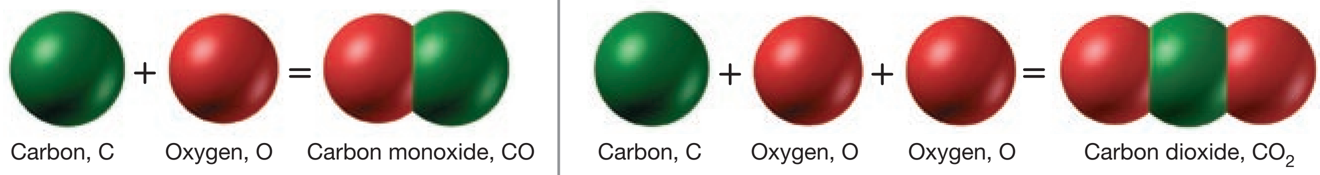
Atoms and the Law of Conservation of Mass

(a) An atom of carbon, C, and an atom of oxygen, O, can combine chemically to form a molecule of carbon monoxide, CO. The mass of the CO molecule is equal to the mass of the C atom plus the mass of the O atom.

(b) The reverse holds true in a reaction in which a CO molecule is broken down into its elements.

FIGURE 1.3

Law of Multiple Proportions



(a) CO molecules are always composed of one C atom and one O atom.

(b) CO_2 molecules are always composed of one C atom and two O atoms. Note that a molecule of carbon dioxide contains twice as many oxygen atoms as does a molecule of carbon monoxide.

Figure 1.3 illustrates how Dalton's atomic theory explained the other laws. The law of definite proportions results from the fact that a given chemical compound always contains the same combinations of atoms. As for the law of multiple proportions, in the case of the carbon oxides, the 2-to-1 ratio of oxygen masses results because carbon dioxide always contains twice as many atoms of oxygen (per atom of carbon) as does carbon monoxide.

MAIN IDEA

Atoms can be subdivided into smaller particles.

By relating atoms to the measurable property of mass, Dalton turned Democritus's *idea* into a *scientific theory* that could be tested by experiment. But not all aspects of Dalton's atomic theory have proven to be correct. For example, today we know that atoms are divisible into even smaller particles (although the law of conservation of mass still holds true for chemical reactions). And, as you will see in Section 3, we know that a given element can have atoms with different masses. Atomic theory has not been discarded—only modified! The important concepts that (1) all matter is composed of atoms and that (2) atoms of any one element differ in properties from atoms of another element remain unchanged.

QuickLAB CONSTRUCTING A MODEL

QUESTION

How can you construct a model of an unknown object by (1) making inferences about an object that is in a closed container and (2) touching the object without seeing it?

PROCEDURE

Record all of your results in a data table.

1. Your teacher will provide you with a can that is covered by a sock sealed with tape. Without unsealing the container, try to determine the number of objects inside the can as well as the mass, shape, size, composition, and texture of each object. To do this, you may carefully tilt or shake the can. Record your observations in a data table.
2. Remove the tape from the top of the sock. Do *not* look inside the can. Put one hand through the opening, and make the same observations as in step 1 by handling the objects. To make more-accurate estimations, practice estimating the sizes and masses of some known objects outside the can. Then compare your estimates of these objects with actual measurements using a metric ruler and a balance.

DISCUSSION

1. Scientists often use more than one method to gather data. How was this illustrated in the investigation?
2. Of the observations you made, which were qualitative and which were quantitative?
3. Using the data you gathered, draw a model of the unknown object(s) and write a brief summary of your conclusions.

MATERIALS

- can covered by a sock sealed with tape
- one or more objects that fit in the container
- metric ruler
- balance

SAFETY

Wear safety goggles and an apron.

SECTION 1 FORMATIVE ASSESSMENT

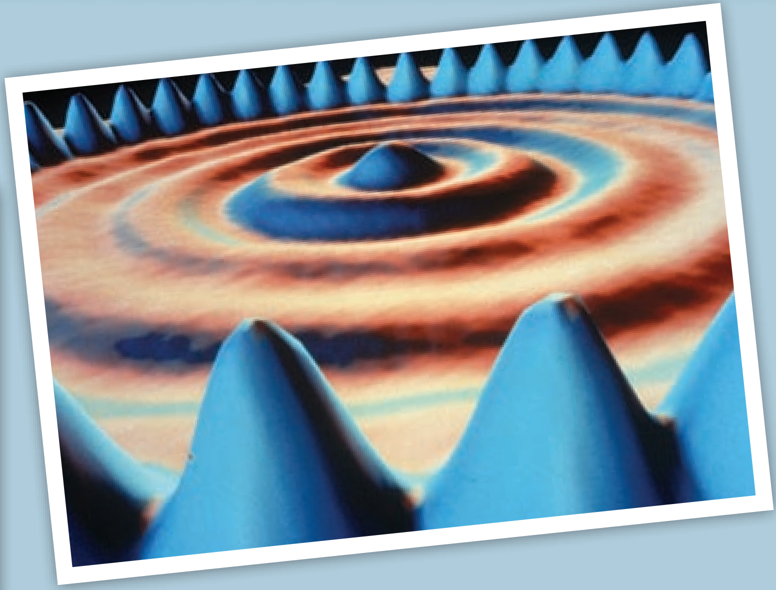
Reviewing Main Ideas

1. List the five main points of Dalton's atomic theory.
2. What chemical laws can be explained by Dalton's theory?

Critical Thinking

3. **ANALYZING INFORMATION** Three compounds containing potassium and oxygen are compared. Analysis shows that for each 1.00 g of O, the compounds have 1.22 g, 2.44 g, and 4.89 g of K, respectively. Show how these data support the law of multiple proportions.

Physical Chemist



This STM image shows a "corral" of iron atoms on a copper surface.

Physical chemists focus on understanding the physical properties of atoms and molecules. They are driven by a curiosity of what makes things work at the level of atoms, and they enjoy being challenged. In addition to chemistry, they study mathematics and physics extensively. Laboratory courses involving experience with electronics and optics are typically part of their training. Often, they enjoy working with instruments and computers. Physical chemists can be experimentalists or theoreticians. They use sophisticated instruments to make measurements or high-powered computers to perform intensive calculations. The instruments used include lasers, electron microscopes, nuclear magnetic resonance spectrometers, mass spectrometers, and particle accelerators. Physical chemists work in industry, government laboratories, research institutes, and academic institutions. Because physical chemists work on a wide range of problems, taking courses in other science disciplines is important.

Scanning Tunneling Microscopy

For years, scientists have yearned for the ability to "see" individual atoms. Because atoms are so small, this had been nothing more than a dream. Now, the scanning tunneling microscope, STM, gives scientists the ability to look at individual atoms. It was invented in 1981 by Gerd Binnig and Heinrich Rohrer, scientists working for IBM in Zurich, Switzerland. They shared the 1986 Nobel Prize in physics for their discovery.

The basic principle of STM is based on the current that exists between a metallic needle that is sharpened to a single atom, the probe, and a conducting sample. As the probe passes above the surface of the sample at a distance of one or two atoms, electrons can "tunnel" from the needle tip to the sample's surface. The probe moves across, or "scans," the surface of the sample. When the probe comes close to the electrons of an individual atom, a signal is produced. A weaker signal is produced between atoms. These signals build a topographical (hill and valley) "map" of conducting and nonconducting regions. The resulting map shows the position and spacing of atoms.

Surface chemistry is a developing subdiscipline in physical chemistry, and STM is an important tool in the field. Scientists use STM to study surface reactions, such as those that take place in catalytic converters. Other areas of research in which STM is useful include semiconductors and microelectronics. Usually, STM is used with materials that conduct, but it has also been used to study biological molecules, such as DNA.

One innovative application of STM is the ability to position individual atoms. The figure shows the result of moving individual atoms. First, iron atoms were placed on a copper surface. Then, individual iron atoms were picked up by the probe and placed in position. The result is a "quantum corral" of 48 iron atoms on the surface of copper. The diameter of the corral is about 14 nm.

Questions

1. In addition to chemistry, what kinds of courses are important for a student interested in a physical chemistry career?
2. What part of an atom is detected by STM?

SECTION 2

Main Ideas

- Atoms contain positive and negative particles.
- Atoms have small, dense, positively-charged nuclei.
- A nucleus contains protons and neutrons.
- The radii of atoms are expressed in picometers.

VIRGINIA STANDARDS

CH.2 The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of:
CH.2.a average atomic mass, mass number, and atomic number.
CH.2.c mass and charge characteristics of subatomic particles.
CH.1.EKS-26; CH.2.EKS-6; CH.2.EKS-16

The Structure of the Atom

Key Terms
atom
nuclear force

Although John Dalton thought atoms were indivisible, investigators in the late 1800s proved otherwise. As scientific advances allowed a deeper exploration of matter, it became clear that atoms are actually composed of smaller particles and that the number and arrangement of these particles within an atom determine that atom's chemical properties. **Therefore, today we define an atom as the smallest particle of an element that retains the chemical properties of that element.**

All atoms consist of two regions. The *nucleus* is a very small region located at the center of an atom. In every atom, the nucleus is made up of at least one positively charged particle called a *proton* and usually one or more neutral particles called *neutrons*. Surrounding the nucleus is a region occupied by negatively charged particles called *electrons*. This region is very large compared with the size of the nucleus. Protons, neutrons, and electrons are referred to as *subatomic particles*.

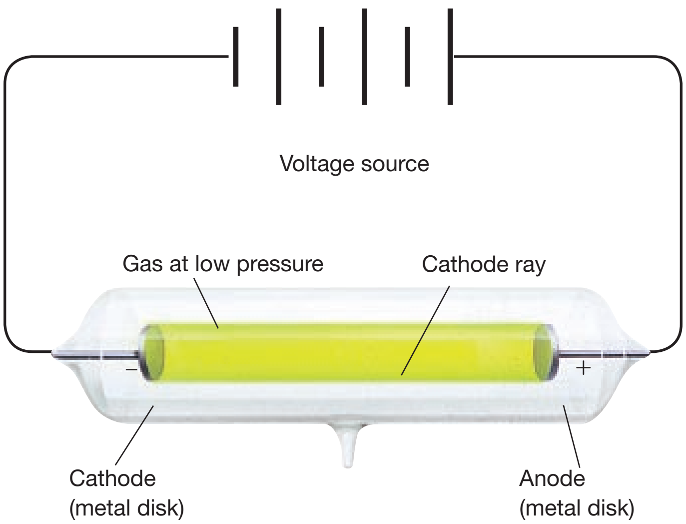
MAIN IDEA

Atoms contain positive and negative particles.

The first discovery of a subatomic particle came in the late 1800s. At that time, many experiments were performed in which electric current was passed through various gases at low pressures. (Gases at atmospheric pressure don't conduct electricity well.) These experiments were carried out in glass tubes, like the one shown in **Figure 2.1,** that had been hooked up to a vacuum pump. Such tubes are known as *cathode-ray tubes.*

FIGURE 2.1

Structure of a Cathode-Ray Tube Particles pass through the tube from the *cathode*, the metal disk connected to the negative terminal of the voltage source, to the *anode*, the metal disk connected to the positive terminal.



Cathode Rays and Electrons

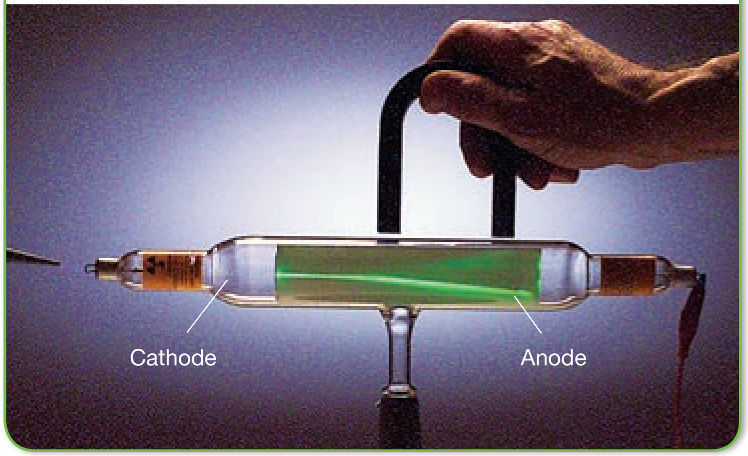
Investigators noticed that when current was passed through the tube, the surface of the tube directly opposite the cathode glowed. They hypothesized that the glow was caused by a stream of particles, which they called a cathode ray. The ray traveled from the cathode to the anode when current was passed through the tube. Experiments devised to test this hypothesis revealed the following observations:

1. Cathode rays were deflected by a magnetic field in the same manner as a wire carrying electric current, which was known to have a negative charge (see **Figure 2.2** on the next page).
2. The rays were deflected away from a negatively charged object.

These observations led to the hypothesis that the particles that compose cathode rays are negatively charged. This hypothesis was strongly supported by a series of experiments carried out in 1897 by the English physicist Joseph John Thomson. In one investigation, he was able to measure the ratio of the charge of cathode-ray particles to their mass. He found that this ratio was always the same, regardless of the metal used to make the cathode or the nature of the gas inside the cathode-ray tube. Thomson concluded that all cathode rays must be composed of identical negatively charged particles, which were named electrons.

FIGURE 2.2

Finding Negative Particles Holding a magnet near a cathode-ray tube (attached to a vacuum pump) causes negatively-charged particles in the beam to be deflected.



Charge and Mass of the Electron

Cathode rays have identical properties regardless of the element used to produce them. Therefore, it was concluded that electrons are present in atoms of all elements. Thus, cathode-ray experiments provided evidence that atoms are divisible and that one of the atom's basic constituents is the negatively charged electron. Thomson's experiment also revealed that the electron has a very large charge-to-mass ratio. In 1909, experiments conducted by the American physicist Robert A. Millikan measured the charge of the electron. Scientists used this information and the charge-to-mass ratio of the electron to determine that the mass of the electron is about one two-thousandth the mass of the simplest type of hydrogen atom, which is the smallest atom known. More-accurate experiments conducted since then indicate that the electron has a mass of 9.109×10^{-31} kg, or 1/1837 the mass of the simplest type of hydrogen atom.

Based on what was learned about electrons, two other inferences were made about atomic structure.

1. Because atoms are electrically neutral, they must contain a positive charge to balance the negative electrons.
2. Because electrons have so much less mass than atoms, atoms must contain other particles that account for most of their mass.

Thomson proposed a model for the atom that is called the *plum pudding model* (after the English dessert). He believed that the negative electrons were spread evenly throughout the positive charge of the rest of the atom. This arrangement is like seeds in a watermelon: the seeds are spread throughout the fruit but do not contribute much to the overall mass. However, shortly thereafter, new experiments disproved this model. Still, the plum pudding model was an important step forward in our modern understanding of the atom, as it represents the first time scientists tried to incorporate the then-revolutionary idea that atoms were not, strictly speaking, indivisible.

CHECK FOR UNDERSTANDING

Analyze Why is it necessary to use experiments such as those of J.J. Thomson and Robert A. Millikan to infer information about electrons?

MAIN IDEA

Atoms have small, dense, positively charged nuclei.

CHECK FOR UNDERSTANDING

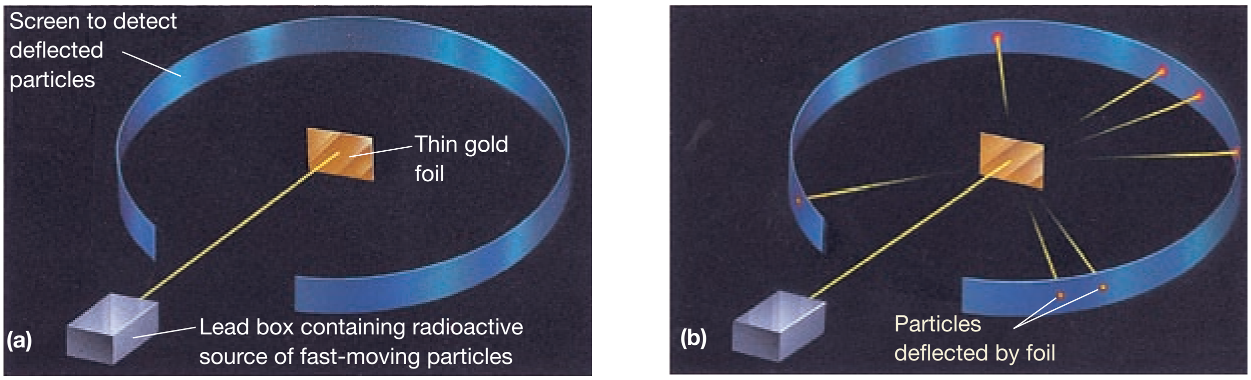
Explain What did Rutherford expect to happen when he fired alpha particles at the gold foil?

More detail of the atom's structure was provided in 1911 by New Zealander Ernest Rutherford and his associates Hans Geiger and Ernest Marsden. The scientists bombarded a thin piece of gold foil with fast-moving *alpha particles,* which are positively charged particles with about four times the mass of a hydrogen atom. Geiger and Marsden assumed that mass and charge were uniformly distributed throughout the atoms of the gold foil, as one would expect from the plum pudding model. They expected the alpha particles to pass through with only a slight deflection, and for the vast majority of the particles, this was the case. However, when the scientists checked for the possibility of wide-angle deflections, they were shocked to find that roughly 1 in 8000 of the alpha particles had actually been deflected back toward the source (see **Figure 2.3**). As Rutherford later exclaimed, it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue paper and it came back and hit you."

After thinking about the startling result for a few months, Rutherford finally came up with an explanation. He reasoned that the deflected alpha particles must have experienced some powerful force within the atom. And he figured that the source of this force must occupy a very small amount of space because so few of the total number of alpha particles had been affected by it. He concluded that the force must be caused by a very densely packed bundle of matter with a positive electric charge. Rutherford called this positive bundle of matter the nucleus (see **Figure 2.4** on the next page).

Rutherford had discovered that the volume of a nucleus was very small compared with the total volume of an atom. In fact, if the nucleus were the size of a marble, then the size of the atom would be about the size of a football field. But where were the electrons? This question was not answered until Rutherford's student, Niels Bohr, proposed a model in which electrons surrounded the positively charged nucleus as the planets surround the sun. Bohr's model is discussed in a later chapter.

FIGURE 2.3



a) Geiger and Marsden bombarded a thin piece of gold foil with a narrow beam of alpha particles.

b) Some of the particles were deflected by the gold foil back toward their source.

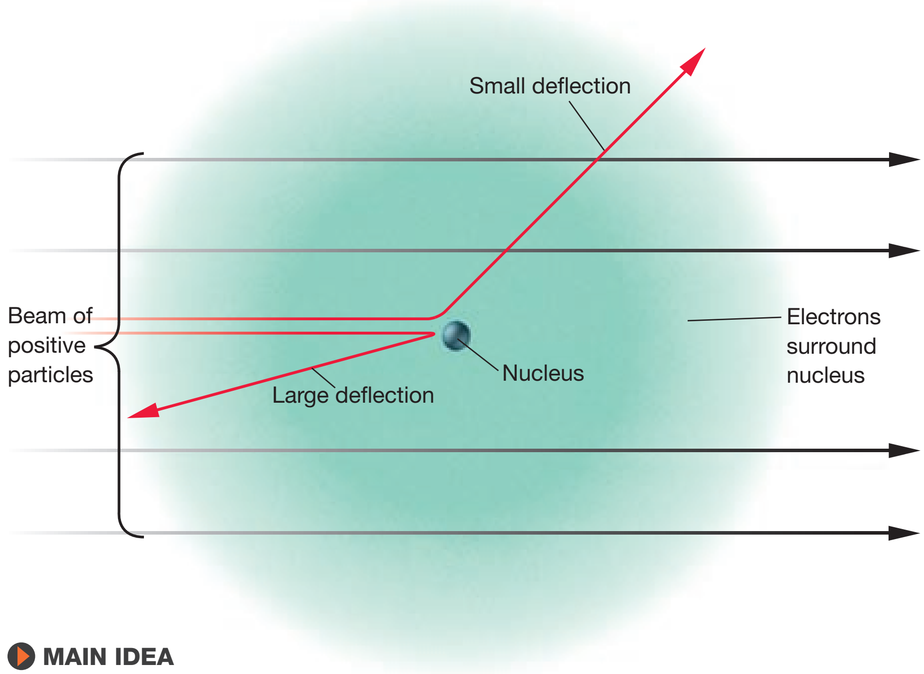


FIGURE 2.4

Finding the Nucleus Rutherford reasoned that each atom in the gold foil contained a small, dense, positively charged nucleus surrounded by electrons. A small number of the alpha particles directed toward the foil were deflected by the tiny nucleus (red arrows). Most of the particles passed through undisturbed (black arrows).

MAIN IDEA

A nucleus contains protons and neutrons.

Except for the nucleus of the simplest type of hydrogen atom (discussed in the next section), all atomic nuclei are made of two kinds of particles, protons and neutrons. A proton has a positive charge equal in magnitude to the negative charge of an electron. Atoms are electrically neutral because they contain equal numbers of protons and electrons. A neutron is electrically neutral.

The simplest hydrogen atom consists of a single-proton nucleus with a single electron moving about it. A proton has a mass of 1.673×10^{-27} kg, which is 1836 times greater than the mass of an electron and 1836/1837, or virtually all, of the mass of the simplest hydrogen atom. All atoms besides the simplest hydrogen atom also have neutrons. The mass of a neutron is 1.675×10^{-27} kg—slightly larger than that of a proton.

The nuclei of atoms of different elements differ in their number of protons and, therefore, in the amount of positive charge they possess. Thus, the number of protons determines that atom's identity. Physicists have identified other subatomic particles, but particles other than electrons, protons, and neutrons have little effect on the chemical properties of matter. **Figure 2.5** on the next page summarizes the properties of electrons, protons, and neutrons.

Forces in the Nucleus

Generally, particles that have the same electric charge repel one another. Therefore, we would expect a nucleus with more than one proton to be unstable. However, when two protons are extremely close to each other, there is a strong attraction between them. In fact, as many as 83 protons can exist close together to help form a stable nucleus. A similar attraction exists when neutrons are very close to each other or when protons and neutrons are very close together. **These short-range proton-neutron, proton-proton, and neutron-neutron forces hold the nuclear particles together and are referred to as nuclear forces.**

FIGURE 2.5

PROPERTIES OF SUBATOMIC PARTICLES					
Particle	Symbols	Relative electric charge	Mass number	Relative mass (u*)	Actual mass (kg)
Electron	e^-, ${}^{0}_{-1}e$	−1	0	0.000 5486	9.109×10^{-31}
Proton	p^+, ${}^{1}_{1}H$	+1	1	1.007 276	1.673×10^{-27}
Neutron	n°, ${}^{1}_{0}n$	0	1	1.008 665	1.675×10^{-27}

*1 u (unified atomic mass unit) = $1.660\ 5402 \times 10^{-27}$ kg

MAIN IDEA

The radii of atoms are expressed in picometers.

It is convenient to think of the region occupied by the electrons as an electron cloud—a cloud of negative charge. The radius of an atom is the distance from the center of the nucleus to the outer portion of this electron cloud. Because atomic radii are so small, they are expressed using a unit that is more convenient for the sizes of atoms. This unit is the picometer. The abbreviation for the picometer is *pm* (1 pm = 10^{-12} m = 10^{-10} cm). To get an idea of how small a picometer is, consider that 1 cm is the same fractional part of 10^3 km (about 600 mi) as 100 pm is of 1 cm. Atomic radii range from about 40 to 270 pm. By contrast, the nuclei of atoms have much smaller radii, about 0.001 pm. Nuclei also have incredibly high densities, about 2×10^8 metric tons/cm^3.

SECTION 2 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. Define each of the following:
 a. atom
 b. electron
 c. nucleus
 d. proton
 e. neutron

2. Describe one conclusion made by each of the following scientists that led to the development of the current atomic theory:
 a. Thomson
 b. Millikan
 c. Rutherford

3. Compare the three subatomic particles in terms of location in the atom, mass, and relative charge.

4. Why are cathode-ray tubes, like the one in **Figure 2.1**, connected to a vacuum pump?

Critical Thinking

5. EVALUATING IDEAS Nuclear forces are said to hold protons and neutrons together. What is it about the composition of the nucleus that requires the concept of nuclear forces?

SECTION 3

Counting Atoms

Main Ideas

- All atoms of an element must have the same number of protons, but not neutrons.
- Atomic mass is a relative measure.
- Average atomic mass is a weighted value.
- A relative mass scale makes counting atoms possible.

Key Terms

atomic number	nuclide	mole
isotope	unified atomic mass unit	Avogadro's number
mass number	average atomic mass	molar mass

VIRGINIA STANDARDS

CH.2 The student will investigate and understand that the placement of elements on the periodic **table** is a function of their atomic structure. The periodic table is a tool used for the investigations of:
CH.2.a average atomic mass, mass number, and atomic number.
CH.2.b isotopes, half lives, and radioactive decay.
CH.4.a The student will investigate and understand that quantities in a chemical reaction are based on molar relationships. Key concepts include: Avogadro's principle and molar volume.
CH.2.EKS-2; CH.2.EKS-3; CH.4.EKS-1

Consider neon, Ne, the gas used in many illuminated signs. Neon is a minor component of the atmosphere. In fact, dry air contains only about 0.002% neon. And yet there are about 5×10^{17} atoms of neon present in each breath you inhale. In most experiments, atoms are much too small to be measured individually. Chemists can analyze atoms quantitatively, however, by knowing fundamental properties of the atoms of each element. In this section, you will be introduced to some of the basic properties of atoms. You will then discover how to use this information to count the number of atoms of an element in a sample with a known mass. You will also become familiar with the *mole*, a special unit used by chemists to express amounts of particles, such as atoms and molecules.

MAIN IDEA

All atoms of an element must have the same number of protons, but not neutrons.

All atoms contain the same particles. Yet all atoms are not the same. Atoms of different elements have different numbers of protons. Atoms of the same element all have the same number of protons. **The atomic number (Z) of an element is the number of protons of each atom of that element.**

Look at a periodic table. In most, an element's atomic number is indicated above its symbol, and the elements are placed in order of increasing atomic number. Hydrogen, H, is at the upper left of the table and has an atomic number of 1. All atoms of the element hydrogen have one proton. Next in order is helium, He, which has two protons. Lithium, Li, has three protons (see **Figure 3.1**); beryllium, Be, has four protons; and so on. The atomic number identifies an element. If the number of protons in the nucleus of an atom were to change, that atom would become a different element.

FIGURE 3.1

Atomic Numbers The atomic number in this periodic table entry reveals that an atom of lithium has three protons in its nucleus.

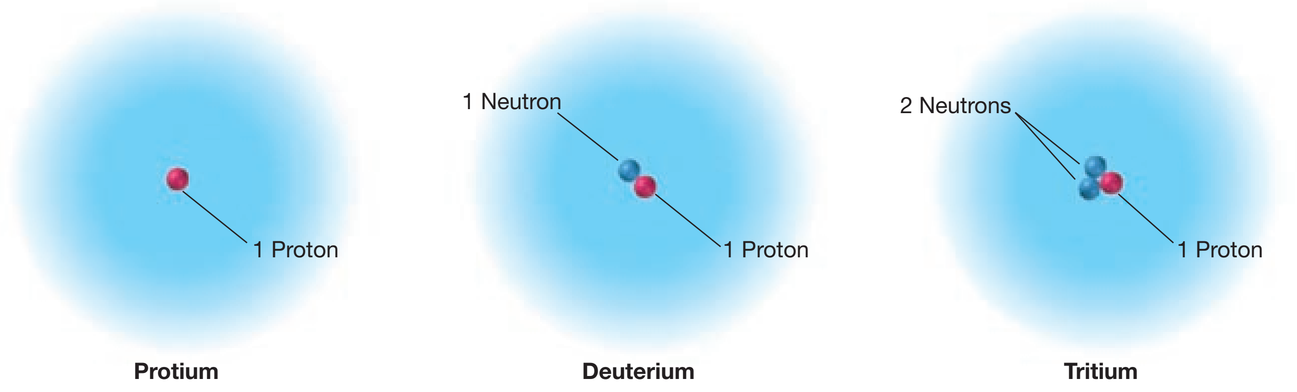
3
Li
Lithium
6.941
[He]$2s^1$

Isotopes

But just because all hydrogen atoms, for example, have only a single proton, it doesn't mean they all have the same number of neutrons, or even any neutrons at all. In fact, three types of hydrogen atoms are known. The most common type of hydrogen is sometimes called *protium*. It accounts for 99.9885% of the hydrogen atoms found on Earth, and its nucleus consists of only a single proton. Another type of hydrogen, *deuterium*, accounts for 0.0115% of Earth's hydrogen atoms; its nucleus has one proton and one neutron. The third form of hydrogen, *tritium*, has one proton and two neutrons in its nucleus. Tritium is radioactive so it is not very common at all on Earth; however, it is still hydrogen.

FIGURE 3.2

Three Isotopes of Hydrogen The nuclei of different isotopes of the same element have the same number of protons but different numbers of neutrons.



Protium, deuterium, and tritium are isotopes of hydrogen. **Isotopes are atoms of the same element that have different masses.** The isotopes of a particular element all have the same number of protons and electrons but different numbers of neutrons. In all three isotopes of hydrogen, shown in **Figure 3.2,** the positive charge of the single proton is balanced by the negative charge of the electron. Most of the elements consist of mixtures of isotopes. Tin has 10 stable isotopes, for example—the most of any element.

The atoms in any sample of an element you may find most likely will be a mixture of several isotopes in various proportions. The detection of these isotopes and determination of their relative proportions has become extremely precise. So precise that scientists can determine where some elements come from by measuring the percentages of different isotopes in a sample.

Mass Number

Identifying an isotope requires knowing both the name or atomic number of the element and the mass of the isotope. **The mass number is the total number of protons and neutrons that make up the nucleus of an isotope.** The three isotopes of hydrogen described earlier have mass numbers 1, 2, and 3, as shown in **Figure 3.3**.

FIGURE 3.3

MASS NUMBERS OF HYDROGEN ISOTOPES

	Atomic number (number of protons)	Number of neutrons	Mass number (protons + neutrons)
Protium	1	0	$1 + 0 = 1$
Deuterium	1	1	$1 + 1 = 2$
Tritium	1	2	$1 + 2 = 3$

Identifying Isotopes

There are two methods for specifying isotopes. In the first, the mass number appears with a hyphen after the name of the element. Tritium, for example, is written as hydrogen-3. We call this method *hyphen notation.* The uranium isotope with mass number 235, commonly used as fuel for nuclear power plants, is known as uranium-235. The second method shows the composition of a nucleus using the isotope's *nuclear symbol.* So uranium-235 is shown as $^{235}_{92}U$. The superscript indicates the mass number (protons + neutrons). The subscript indicates the atomic number (number of protons). The number of neutrons is found by subtracting the atomic number from the mass number.

$$\text{mass number} - \text{atomic number} = \text{number of neutrons}$$
$$235 \text{ (protons + neutrons)} - 92 \text{ protons} = 143 \text{ neutrons}$$

Figure 3.4 gives the names, symbols, and compositions of the isotopes of hydrogen and helium. **Nuclide is a general term for a specific isotope of an element.**

FIGURE 3.4

ISOTOPES OF HYDROGEN AND HELIUM

Isotope	Nuclear symbol	Number of protons	Number of electrons	Number of neutrons
Hydrogen-1 (protium)	$^{1}_{1}H$	1	1	0
Hydrogen-2 (deuterium)	$^{2}_{1}H$	1	1	1
Hydrogen-3 (tritium)	$^{3}_{1}H$	1	1	2
Helium-3	$^{3}_{2}He$	2	2	1
Helium-4	$^{4}_{2}He$	2	2	2

Sub-Atomic Particles

Sample Problem A **How many protons, electrons, and neutrons are there in an atom of chlorine-37?**

1 ANALYZE

Given: name and mass number of chlorine-37

Unknown: numbers of protons, electrons, and neutrons

2 PLAN

atomic number = number of protons = number of electrons
mass number = number of neutrons + number of protons

Continued

3 SOLVE

The mass number of chlorine-37 is 37. Consulting the periodic table reveals that chlorine's atomic number is 17. Therefore we know that

$$\text{atomic number} = \text{number of protons} = \text{number of electrons} = 17 \text{ protons and } 17 \text{ electrons}$$

$$\text{number of neutrons} = \text{mass number} - \text{atomic number} = 37 - 17 = 20 \text{ neutrons}$$

An atom of chlorine-37 is made up of 17 electrons, 17 protons, and 20 neutrons.

4 CHECK YOUR WORK

The number of protons in a neutral atom equals the number of electrons. The sum of the protons and neutrons equals the given mass number ($17 + 20 = 37$).

Practice Answers in Appendix E

1. How many protons, electrons, and neutrons make up an atom of bromine-80?
2. Write the nuclear symbol for carbon-13.
3. Write the hyphen notation for the isotope with 15 electrons and 15 neutrons.

MAIN IDEA

Atomic mass is a relative measure.

Masses of atoms expressed in grams are very small. As we shall see, an atom of oxygen-16, for example, has a mass of 2.656×10^{-23} g. For most chemical calculations it is more convenient to use *relative* atomic masses. As you learned when you studied scientific measurement, scientists use standards of measurement that are constant and are the same everywhere. In order to set up a relative scale of atomic mass, one atom has been arbitrarily chosen as the standard and assigned a mass value. The masses of all other atoms are expressed in relation to this standard.

The standard used by scientists to compare units of atomic mass is the carbon-12 atom, which has been arbitrarily assigned a mass of exactly 12 unified atomic mass units, or 12 u. **One unified atomic mass unit, or 1 u, is exactly 1/12 the mass of a carbon-12 atom.** The atomic mass of any other atom is determined by comparing it with the mass of the carbon-12 atom. The hydrogen-1 atom has an atomic mass of *about* 1/12 that of the carbon-12 atom, or about 1 u. The precise value of the atomic mass of a hydrogen-1 atom is 1.007 825 u. An oxygen-16 atom has about 16/12 (or 4/3) the mass of a carbon-12 atom. Careful measurements show the atomic mass of oxygen-16 to be 15.994 915 u. The mass of a magnesium-24 atom is found to be slightly less than twice that of a carbon-12 atom. Its atomic mass is 23.985 042 u.

Some additional examples of the atomic masses of the naturally occurring isotopes of several elements are given in **Figure 3.5** on the next page. Isotopes of an element may occur naturally, or they may be made in the laboratory (*artificial isotopes*). *Although isotopes have different masses, they do not differ significantly in their chemical behavior.*

The masses of subatomic particles can also be expressed on the atomic mass scale (see **Figure 2.5**). The mass of the electron is 0.000 548 6 u, that of the proton is 1.007 276 u, and that of the neutron is 1.008 665 u. Note that the proton and neutron masses are close, but not equal, to 1 u. You have learned that the mass number is the total number of protons and neutrons that make up the nucleus of an atom. You can now see that the mass number and relative atomic mass of a given nuclide are quite close to each other. They are not identical, because the proton and neutron masses deviate slightly from 1 u and the atomic masses include electrons. Also, as you will read in a later chapter, a small amount of mass is changed to energy in the creation of a nucleus from its protons and neutrons.

Chemistry EXPLORERS

Discovery of Element 43

The discovery of element 43, technetium, is credited to Carlo Perrier and Emilio Segrè, who artificially produced it in 1937. However, scientists have found minute traces of technetium in the Earth's crust that result from the fission of uranium. Astronomers have also discovered technetium in S-type stars.

MAIN IDEA

Average atomic mass is a weighted value.

Most elements occur naturally as mixtures of isotopes, as indicated in **Figure 3.5** (see next page). Scientists determine the average mass of a sample of an element's isotopes by determining the percentages of each of the isotopes and then giving the proper weight to each value. **Average atomic mass** **is the weighted average of the atomic masses of the naturally occurring isotopes of an element.** Unlike atomic number, average atomic mass is a statistical calculation. Different samples of the same element can differ in their relative abundance of isotopes.

The following is an example of how to calculate a *weighted average*. Suppose you have a box containing two sizes of marbles. If 25% of the marbles have masses of 2.00 g each and 75% have masses of 3.00 g each, how is the weighted average calculated? You could count the number of each type of marble, calculate the total mass of the mixture, and divide by the total number of marbles. If you had 100 marbles, the calculations would be as follows:

$$\text{25 marbles} \times 2.00\ \text{g} = 50\ \text{g}$$
$$\text{75 marbles} \times 3.00\ \text{g} = 225\ \text{g}$$

Adding these masses gives the total mass of the marbles.

$$50\ \text{g} + 225\ \text{g} = 275\ \text{g}$$

Dividing the total mass by 100 gives an average marble mass of 2.75 g.

A simpler method is to multiply the mass of each marble by the decimal fraction representing its percentage in the mixture. Then add the products.

$$25\% = 0.25 \qquad 75\% = 0.75$$
$$(2.00\ \text{g} \times 0.25) + (3.00\ \text{g} \times 0.75) = 2.75\ \text{g}$$

FIGURE 3.5

ATOMIC MASSES AND ABUNDANCES OF SEVERAL NATURALLY OCCURRING ISOTOPES

Isotope	Mass number	Percentage natural abundance	Atomic mass (u)	Average atomic mass of element (u)
Hydrogen-1	1	99.9885	1.007 825	1.007 94
Hydrogen-2	2	0.0115	2.014 102	
Carbon-12	12	98.93	12 (by definition)	
Carbon-13	13	1.07	13.003 355	12.0107
Oxygen-16	16	99.757	15.994 915	15.9994
Oxygen-17	17	0.038	16.999 132	
Oxygen-18	18	0.205	17.999 160	
Copper-63	63	69.15	62.929 601	63.546
Copper-65	65	30.85	64.927 794	
Cesium-133	133	100	132.905 447	132.905
Uranium-234	234	0.0054	234.040 945	
Uranium-235	235	0.7204	235.043 922	238.029
Uranium-238	238	99.2742	238.050 784	

Calculating Average Atomic Mass

The average atomic mass of an element depends on both the mass and the relative abundance of each of the element's isotopes. For example, naturally occurring copper consists of 69.15% copper-63, which has an atomic mass of 62.929 601 u, and 30.85% copper-65, which has an atomic mass of 64.927 794 u. The average atomic mass of copper can be calculated by multiplying the atomic mass of each isotope by its relative abundance (expressed in decimal form) and adding the results.

$$0.6915 \times 62.929\,601\ \text{u} + 0.3085 \times 64.927\,794\ \text{u} = 63.55\ \text{u}$$

The calculated average atomic mass of naturally occurring copper is 63.55 u.

The average atomic mass is included for the elements listed in **Figure 3.5**. As illustrated in the table, most atomic masses are known to four or more significant figures. *In this book, an element's atomic mass is usually rounded to two decimal places before it is used in a calculation.*

MAIN IDEA

A relative mass scale makes counting atoms possible.

The relative atomic mass scale makes it possible to know how many atoms of an element are present in a sample of the element with a measurable mass. Three very important concepts—the mole, Avogadro's number, and molar mass—provide the basis for relating masses in grams to numbers of atoms.

PREMIUM CONTENT
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Avogadro's Number

The Mole

The mole is the SI unit for amount of substance. **A mole (abbreviated mol) is the amount of a substance that contains as many particles as there are atoms in exactly 12 g of carbon-12.** The mole is a counting unit, just like a dozen is. We don't usually buy 12 or 24 ears of corn; we order one dozen or two dozen. Similarly, a chemist may want 1 mol of carbon, or 2 mol of iron, or 2.567 mol of calcium. In the sections that follow, you will see how the mole relates to masses of atoms and compounds.

Avogadro's Number

The number of particles in a mole has been experimentally determined in a number of ways. The best modern value is $6.022\,141\,79 \times 10^{23}$. This means that exactly 12 g of carbon-12 contains $6.022\,141\,79 \times 10^{23}$ carbon-12 atoms.

The number of particles in a mole is known as Avogadro's number, named for the nineteenth-century Italian scientist Amedeo Avogadro, whose ideas were crucial in explaining the relationship between mass and numbers of atoms. **Avogadro's number—$6.022\,141\,79 \times 10^{23}$—is the number of particles in exactly one mole of a pure substance.** For most purposes, Avogadro's number is rounded to 6.022×10^{23}.

To get a sense of how large Avogadro's number is, consider the following: If every person living on Earth (6.8 billion people) worked to count the atoms in one mole of an element, and if each person counted continuously at a rate of one atom per second, it would take about 3 million years for all the atoms to be counted.

Molar Mass

An alternative definition of *mole* is the amount of a substance that contains Avogadro's number of particles. Can you calculate the approximate mass of one mole of helium atoms? You know that a mole of carbon-12 atoms has a mass of exactly 12 g and that a carbon-12 atom has an atomic mass of 12 u. The atomic mass of a helium atom is 4.00 u, which is about one-third the mass of a carbon-12 atom. It follows that a mole of helium atoms will have about one-third the mass of a mole of carbon-12 atoms. Thus, one mole of helium has a mass of about 4.00 g.

The mass of one mole of a pure substance is called the molar mass of that substance. Molar mass is usually written in units of g/mol. The molar mass of an element is numerically equal to the atomic mass of the element in unified atomic mass units (which can be found in the periodic table). For example, the molar mass of lithium, Li, is 6.94 g/mol, while the molar mass of mercury, Hg, is 200.59 g/mol (rounding each value to two decimal places). The molar mass of an element contains one mole of atoms. For example, 4.00 g of helium, 6.94 g of lithium, and 200.59 g of mercury all contain a mole of atoms. **Figure 3.6** shows molar masses of three common elements.

FIGURE 3.6

Molar Mass Shown is approximately one molar mass of each of three elements.

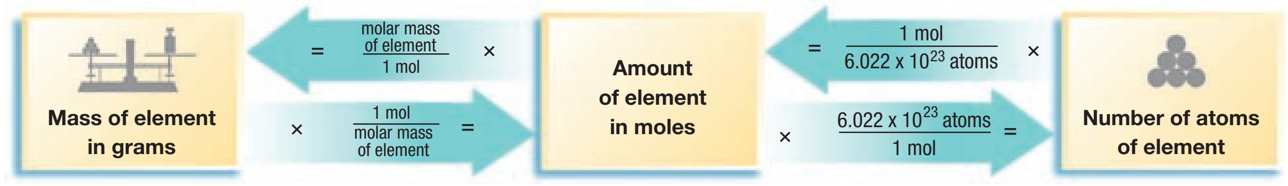
(a) carbon (graphite)

(b) iron (nails)

(c) copper (wire)

FIGURE 3.7

Relating Mass to the Number of Atoms The diagram shows the relationship between mass in grams, amount in moles, and number of atoms of an element in a sample.



Gram/Mole Conversions

Chemists use molar mass as a conversion factor in chemical calculations. For example, the molar mass of helium is 4.00 g He/mol He. To find how many grams of helium there are in two moles of helium, multiply by the molar mass.

$$2.00\text{ mol He} \times \frac{4.00\text{ g He}}{1\text{ mol He}} = 8.00\text{ g He}$$

Figure 3.7 shows how to use molar mass, moles, and Avogadro's number to relate mass in grams, amount in moles, and number of atoms of an element.

Sub-Atomic Particles

Sample Problem B **What is the mass in grams of 3.50 mol of the element copper, Cu?**

1 ANALYZE

Given: 3.50 mol Cu

Unknown: mass of Cu in grams

2 PLAN

amount of Cu in moles ⟶ mass of Cu in grams

According to **Figure 3.7,** the mass of an element in grams can be calculated by multiplying the amount of the element in moles by the element's molar mass.

$$\text{moles Cu} \times \frac{\text{grams Cu}}{\text{moles Cu}} = \text{grams Cu}$$

3 SOLVE

The molar mass of copper from the periodic table is rounded to 63.55 g/mol.

$$3.50\ \cancel{\text{mol Cu}} \times \frac{63.55\text{ g Cu}}{1\ \cancel{\text{mol Cu}}} = 222\text{ g Cu}$$

4 CHECK YOUR WORK

Because the amount of copper in moles was given to three significant figures, the answer was rounded to three significant figures. The size of the answer is reasonable because it is somewhat more than 3.5 times 60.

Continued

Practice Answers in Appendix E

1. What is the mass in grams of 2.25 mol of the element iron, Fe?
2. What is the mass in grams of 0.375 mol of the element potassium, K?
3. What is the mass in grams of 0.0135 mol of the element sodium, Na?
4. What is the mass in grams of 16.3 mol of the element nickel, Ni?

Gram/Mole Conversions

Sample Problem C **A chemist produced 11.9 g of aluminum, Al. How many moles of aluminum were produced?**

1 ANALYZE

Given: 11.9 g Al

Unknown: amount of Al in moles

2 PLAN

mass of Al in grams $\longrightarrow$ amount of Al in moles

As shown in **Figure 3.7**, amount in moles can be obtained by *dividing* mass in grams by molar mass, which is mathematically the same as *multiplying* mass in grams by the *reciprocal* of molar mass.

$$\text{grams Al} \times \frac{\text{moles Al}}{\text{grams Al}} = \text{moles Al}$$

3 SOLVE

The molar mass of aluminum from the periodic table is rounded to 26.98 g/mol.

$$11.9\ \cancel{\text{g Al}} \times \frac{1\ \text{mol Al}}{26.98\ \cancel{\text{g Al}}} = 0.441\ \text{mol Al}$$

4 CHECK YOUR WORK

The answer is correctly given to three significant figures. The answer is reasonable because 11.9 g is somewhat less than half of 26.98 g.

Practice Answers in Appendix E

1. How many moles of calcium, Ca, are in 5.00 g of calcium?
2. How many moles of gold, Au, are in 3.60×10^{-5} g of gold?
3. How many moles of zinc, Zn, are in 0.535 g of zinc?

CHECK FOR UNDERSTANDING

Describe Although using the mole unit of measurement is often preferable for doing calculations to using Avogadro's number, at other times, using Avogadro's number is necessary. Describe a circumstance under which this is true.

Conversions with Avogadro's Number

Avogadro's number can be used to find the number of atoms of an element from the amount in moles or to find the amount of an element in moles from the number of atoms. While these types of problems are less common in chemistry than converting between amount in moles and mass in grams, they are useful in demonstrating the meaning of Avogadro's number. Note that in these calculations, Avogadro's number is expressed in units of atoms per mole.

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Conversions with Avogadro's Number

Sample Problem D **How many moles of silver, Ag, are in 3.01×10^{23} atoms of silver?**

1 ANALYZE

Given: 3.01×10^{23} atoms of Ag

Unknown: amount of Ag in moles

2 PLAN

number of atoms of Ag $\longrightarrow$ amount of Ag in moles

From **Figure 3.7**, we know that number of atoms is converted to amount in moles by dividing by Avogadro's number. This is equivalent to multiplying numbers of atoms by the reciprocal of Avogadro's number.

$$\text{Ag atoms} \times \frac{\text{moles Ag}}{\text{Avogadro's number of Ag atoms}} = \text{moles Ag}$$

3 SOLVE

$$3.01 \times 10^{23}\ \cancel{\text{Ag atoms}} \times \frac{1\ \text{mol Ag}}{6.022 \times 10^{23}\ \cancel{\text{Ag atoms}}} = 0.500\ \text{mol Ag}$$

4 CHECK YOUR WORK

The answer is correct—units cancel correctly and the number of atoms is one-half of Avogadro's number.

Practice Answers in Appendix E

1. How many moles of lead, Pb, are in 1.50×10^{12} atoms of lead?
2. How many moles of tin, Sn, are in 2500 atoms of tin?
3. How many atoms of aluminum, Al, are in 2.75 mol of aluminum?

Conversions with Avogadro's Number

Sample Problem E **What is the mass in grams of 1.20×10^{8} atoms of copper, Cu?**

1 ANALYZE

Given: 1.20×10^{8} atoms of Cu

Unknown: mass of Cu in grams

Continued

2 PLAN

$$\text{number of atoms of Cu} \longrightarrow \text{amount of Cu in moles} \longrightarrow \text{mass of Cu in grams}$$

As indicated in **Figure 3.7,** the given number of atoms must first be converted to amount in moles by dividing by Avogadro's number. Amount in moles is then multiplied by molar mass to yield mass in grams.

$$\text{Cu atoms} \times \frac{\text{moles Cu}}{\text{Avogadro's number of Cu atoms}} \times \frac{\text{grams Cu}}{\text{moles Cu}} = \text{grams Cu}$$

3 SOLVE

The molar mass of copper from the periodic table is rounded to 63.55 g/mol.

$$1.20 \times 10^{8}\ \text{Cu}\ \cancel{\text{atoms}} \times \frac{1\ \cancel{\text{mol Cu}}}{6.022 \times 10^{23}\ \cancel{\text{Cu atoms}}} \times \frac{63.55\ \text{g Cu}}{1\ \cancel{\text{mol Cu}}} = 1.27 \times 10^{-14}\ \text{g Cu}$$

4 CHECK YOUR WORK

Units cancel correctly to give the answer in grams. The size of the answer is reasonable—108 has been divided by about 1024 and multiplied by about 102.

Practice

1. What is the mass in grams of 7.5×10^{15} atoms of nickel, Ni?
2. How many atoms of sulfur, S, are in 4.00 g of sulfur?
3. What mass of gold, Au, contains the same number of atoms as 9.0 g of aluminum, Al?

SECTION 3 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. Explain each of the following:
   a. atomic number
   b. mass number
   c. relative atomic mass
   d. average atomic mass
   e. mole
   f. Avogadro's number
   g. molar mass
   h. isotope
2. Determine the number of protons, electrons, and neutrons in each of the following isotopes:
   a. sodium-23
   b. calcium-40
   c. $^{64}_{22}\text{Cu}$
   d. $^{108}_{47}\text{Ag}$
3. Write the nuclear symbol and hyphen notation for each of the following isotopes:
   a. mass number of 28 and atomic number of 14
   b. 26 protons and 30 neutrons
4. To two decimal places, what is the relative atomic mass and the molar mass of the element potassium, K?
5. Determine the mass in grams of the following:
   a. 2.00 mol N
   b. 3.01×10^{23} atoms Cl
6. Determine the amount in moles of the following:
   a. 12.15 g Mg
   b. 1.50×10^{23} atoms F

Critical Thinking

7. **ANALYZING DATA** Beaker A contains 2.06 mol of copper, and Beaker B contains 222 grams of silver. Which beaker contains the larger mass? Which beaker has the larger number of atoms?

Math Tutor

Conversion Factors

Most calculations in chemistry require that all measurements of the same quantity (mass, length, volume, temperature, and so on) be expressed in the same unit. To change the units of a quantity, you can multiply the quantity by a conversion factor. With SI units, such conversions are easy because units of the same quantity are related by multiples of 10, 100, 1000, or 1 million. Suppose you want to convert a given amount in milliliters to liters. You can use the relationship 1 L = 1000 mL. From this relationship, you can derive the following conversion factors.

$$\frac{1000 \text{ mL}}{1 \text{ L}} \text{ and } \frac{1 \text{ L}}{1000 \text{ mL}}$$

The correct strategy is to multiply the given amount (in mL) by the conversion factor that allows milliliter units to cancel out and liter units to remain. Using the second conversion factor will give you the units you want.

These conversion factors are based on an exact definition (1000 mL = 1 L exactly), so significant figures do not apply to these factors. The number of significant figures in a converted measurement depends on the certainty of the measurement you start with.

Sample Problem

A sample of aluminum has a mass of 0.087 g. What is the sample's mass in milligrams?

Based on SI prefixes, you know that 1 g = 1000 mg. Therefore, the possible conversion factors are

$$\frac{1000 \text{ mg}}{1 \text{ g}} \text{ and } \frac{1 \text{ g}}{1000 \text{ mg}}$$

The first conversion factor cancels grams, leaving milligrams.

$$0.087 \cancel{\text{g}} \times \frac{1000 \text{ mg}}{1 \cancel{\text{g}}} = 87 \text{ mg}$$

Notice that the values 0.087 g and 87 mg each have two significant figures.

A sample of a mineral has 4.08×10^{-5} mol of vanadium per kilogram of mass. How many micromoles of vanadium per kilogram does the mineral contain?

The prefix *micro-* specifies $\frac{1}{1\,000\,000}$ or 1×10^{-6} of the base unit.
So, 1 μmol = 1×10^{-6} mol. The possible conversion factors are

$$\frac{1 \ \mu\text{mol}}{1 \times 10^{-6} \text{ mol}} \text{ and } \frac{1 \times 10^{-6} \text{ mol}}{1 \ \mu\text{mol}}$$

The first conversion factor will allow moles to cancel and micromoles to remain.

$$4.08 \times 10^{-5} \cancel{\text{mol}} \times \frac{1 \ \mu\text{mol}}{1 \times 10^{-6} \cancel{\text{mol}}} = 40.8 \ \mu\text{mol}$$

Notice that the values 4.08×10^{-5} mol and 40.8 μmol each have three significant figures.

Practice

1. Express each of the following measurements in the units indicated.
   a. 2250 mg in grams
   b. 59.3 kL in liters
2. Use scientific notation to express each of the following measurements in the units indicated.
   a. 0.000 072 g in micrograms
   b. 3.98×10^{6} m in kilometers

CHAPTER 3 **Summary**

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Concept Maps

SECTION 1 The Atom: From Philosophical Idea to Scientific Theory

- The idea of atoms has been around since the time of the ancient Greeks. In the nineteenth century, John Dalton proposed a scientific theory of atoms that can still be used to explain properties of most chemicals today.
- Matter and its mass cannot be created or destroyed in chemical reactions.
- The mass ratios of the elements that make up a given compound are always the same, regardless of how much of the compound there is or how it was formed.
- If two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element can be expressed as a ratio of small whole numbers.

KEY TERMS

law of conservation of mass
law of definite proportions
law of multiple proportions

SECTION 2 The Structure of the Atom

- Cathode-ray tubes supplied evidence of the existence of electrons, which are negatively charged subatomic particles that have relatively little mass.
- Rutherford found evidence for the existence of the atomic nucleus by bombarding gold foil with a beam of positively charged particles.
- Atomic nuclei are composed of protons, which have an electric charge of $+1$, and (in all but one case) neutrons, which have no electric charge.
- Atomic nuclei have radii of about 0.001 pm (pm = picometers; 1 pm $\times$ 10^{-12} m), and atoms have radii of about 40–270 pm.

KEY TERMS

atom
nuclear forces

SECTION 3 Counting Atoms

- The atomic number of an element is equal to the number of protons of an atom of that element.
- The mass number is equal to the total number of protons and neutrons that make up the nucleus of an atom of that element.
- The unified atomic mass unit (u) is based on the carbon-12 atom and is a convenient unit for measuring the mass of atoms. It equals $1.660\ 540 \times 10^{-24}$ g.
- The average atomic mass of an element is found by calculating the weighted average of the atomic masses of the naturally occurring isotopes of the element.
- Avogadro's number is equal to approximately 6.022×10^{23}. A sample that contains a number of particles equal to Avogadro's number contains a mole of those particles.

KEY TERMS

atomic number
isotope
mass number
nuclide
unified atomic mass unit
average atomic mass
mole
Avogadro's number
molar mass

CHAPTER 3 Review

SECTION 1

The Atom: From Philosophical Idea to Scientific Theory

REVIEWING MAIN IDEAS

1. Explain each of the following in terms of Dalton's atomic theory:
   a. the law of conservation of mass
   b. the law of definite proportions
   c. the law of multiple proportions

2. According to the law of conservation of mass, if element A has an atomic mass of 2 mass units and element B has an atomic mass of 3 mass units, what mass would be expected for compound AB? for compound A_2B_3?

SECTION 2

The Structure of the Atom

REVIEWING MAIN IDEAS

3. a. What is an atom?
   b. What two regions make up all atoms?

4. Describe at least four properties of electrons that were determined based on the experiments of Thomson and Millikan.

5. Summarize Rutherford's model of the atom, and explain how he developed this model based on the results of his famous gold-foil experiment.

6. What number uniquely identifies an element?

SECTION 3

Counting Atoms

REVIEWING MAIN IDEAS

7. a. What are isotopes?
   b. How are the isotopes of a particular element alike?
   c. How are they different?

8. Copy and complete the following table concerning the three isotopes of silicon, Si. (Hint: See Sample Problem A.)

Isotope	Number of protons	Number of electrons	number of neutrons
Si-28			
Si-29			
Si-30			

9. a. What is the atomic number of an element?
   b. What is the mass number of an isotope?
   c. In the nuclear symbol for deuterium, $^{2}_{1}H$, identify the atomic number and the mass number.

10. What is a nuclide?

11. Use the periodic table and the information that follows to write the hyphen notation for each isotope described.
    a. atomic number = 2, mass number = 4
    b. atomic number = 8, mass number = 16
    c. atomic number = 19, mass number = 39

12. a. What nuclide is used as the standard in the relative scale for atomic masses?
    b. What is its assigned atomic mass?

13. What is the atomic mass of an atom if its mass is approximately equal to the following?
    a. $\frac{1}{3}$ that of carbon-12
    b. 4.5 times as much as carbon-12

14. a. What is the definition of a *mole*?
    b. What is the abbreviation for *mole*?
    c. How many particles are in one mole?
    d. What name is given to the number of particles in a mole?

15. a. What is the molar mass of an element?
    b. To two decimal places, write the molar masses of carbon, neon, iron, and uranium.

16. Suppose you have a sample of an element.
    a. How is the mass in grams of the element converted to amount in moles?
    b. How is the mass in grams of the element converted to number of atoms?

PRACTICE PROBLEMS

17. What is the mass in grams of each of the following? (Hint: See Sample Problems B and E.)
- **a.** 1.00 mol Li
- **b.** 1.00 mol Al
- **c.** 1.00 molar mass Ca
- **d.** 1.00 molar mass Fe
- **e.** 6.022×10^{23} atoms C
- **f.** 6.022×10^{23} atoms Ag

18. How many moles of atoms are there in each of the following? (Hint: See Sample Problems C and D.)
- **a.** 6.022×10^{23} atoms Ne
- **b.** 3.011×10^{23} atoms Mg
- **c.** 3.25×10^{5} g Pb
- **d.** 4.50×10^{-12} g O

19. Three isotopes of argon occur in nature—$^{36}_{18}Ar$, $^{38}_{18}Ar$, and $^{40}_{18}Ar$. Calculate the average atomic mass of argon to two decimal places, given the following relative atomic masses and abundances of each of the isotopes: argon-36 (35.97 u; 0.337%), argon-38 (37.96 u; 0.063%), and argon-40 (39.96 u; 99.600%).

20. Naturally occurring boron is 80.20% boron-11 (atomic mass = 11.01 u) and 19.80% of some other isotopic form of boron. What must the atomic mass of this second isotope be in order to account for the 10.81 u average atomic mass of boron? (Write the answer to two decimal places.)

21. How many atoms are there in each of the following?
- **a.** 1.50 mol Na
- **b.** 6.755 mol Pb
- **c.** 7.02 g Si

22. What is the mass in grams of each of the following?
- **a.** 3.011×10^{23} atoms F
- **b.** 1.50×10^{23} atoms Mg
- **c.** 4.50×10^{12} atoms Cl
- **d.** 8.42×10^{18} atoms Br
- **e.** 25 atoms W
- **f.** 1 atom Au

23. Determine the number of atoms in each of the following:
- **a.** 5.40 g B
- **b.** 0.250 mol S
- **c.** 0.0384 mol K
- **d.** 0.025 50 g Pt
- **e.** 1.00×10^{-10} g Au

Mixed Review

REVIEWING MAIN IDEAS

24. Determine the mass in grams of each of the following:
- **a.** 3.00 mol Al
- **b.** 2.56×10^{24} atoms Li
- **c.** 1.38 mol N
- **d.** 4.86×10^{24} atoms Au
- **e.** 6.50 mol Cu
- **f.** 2.57×10^{8} mol S
- **g.** 1.05×10^{18} atoms Hg

25. Copy and complete the following table concerning the properties of subatomic particles.

Particle	Symbol	Mass number	Actual mass	Relative charge
Electron				
Proton				
Neutron				

26. **a.** How is a unified atomic mass unit (u) related to the mass of one carbon-12 atom?
b. What is the relative atomic mass of an atom?

27. **a.** What is the nucleus of an atom?
b. Who is credited with the discovery of the atomic nucleus?
c. Identify the two kinds of particles that make up the nucleus.

28. How many moles of atoms are there in each of the following?
- **a.** 40.1 g Ca
- **b.** 11.5 g Na
- **c.** 5.87 g Ni
- **d.** 150 g S
- **e.** 2.65 g Fe
- **f.** 0.007 50 g Ag
- **g.** 2.25×10^{25} atoms Zn
- **h.** 50 atoms Ba

29. State the law of multiple proportions, and give an example of two compounds that illustrate the law.

30. What is the approximate atomic mass of an atom if its mass is
- **a.** 12 times that of carbon-12?
- **b.** $\frac{1}{2}$ that of carbon-12?

31. What is an electron?

CRITICAL THINKING

32. **Organizing Ideas** Using two chemical compounds as an example, describe the difference between the law of definite proportions and the law of multiple proportions.

33. **Constructing Models** As described in Section 2, the structure of the atom was determined from observations made in painstaking experimental research. Suppose a series of experiments revealed that when an electric current is passed through gas at low pressure, the surface of the cathode-ray tube opposite the anode glows. In addition, a paddle wheel placed in the tube rolls from the anode toward the cathode when the current is on.
    a. In which direction do particles pass through the gas?
    b. What charge do the particles possess?

34. **Analyzing Data** Osmium is the element with the greatest density, 22.58 g/cm^3. How does the density of osmium compare to the density of a typical nucleus of 2×10^8 metric tons/cm^3? (1 metric ton = 1000 kg)

USING THE HANDBOOK

35. Group 14 of the *Elements Handbook* (Appendix A) describes the reactions that produce CO and CO_2. Review this section to answer the following:
    a. When a fuel burns, what determines whether CO or CO_2 will be produced?
    b. What happens in the body if hemoglobin picks up CO?
    c. Why is CO poisoning most likely to occur in homes that are well sealed during cold winter months?

RESEARCH AND WRITING

36. Prepare a report on the series of experiments conducted by Sir James Chadwick that led to the discovery of the neutron.

37. Write a report on the contributions of Amedeo Avogadro that led to the determination of the value of Avogadro's number.

38. Trace the development of the electron microscope, and cite some of its many uses.

39. The study of atomic structure and the nucleus produced a new field of medicine called *nuclear medicine*. Describe the use of radioactive tracers to detect and treat diseases.

ALTERNATIVE ASSESSMENT

40. Observe a cathode-ray tube in operation, and write a description of your observations.

41. **Performance Assessment** Using colored clay, build a model of the nucleus of each of carbon's three naturally occurring isotopes: carbon-12, carbon-13, and carbon-14. Specify the number of electrons that would surround each nucleus.

Standards-Based Assessment

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. A chemical compound always has the same elements in the same proportions by mass regardless of the source of the compound. This is a statement of
A. the law of multiple proportions.
B. the law of isotopes.
C. the law of definite proportions.
D. the law of conservation of mass.

2. An important result of Rutherford's experiments with gold foil was to establish that
A. atoms have mass.
B. electrons have a negative charge.
C. neutrons are uncharged particles.
D. the atom is mostly empty space.

3. Which subatomic particle has a charge of +1?
A. electron
B. neutron
C. proton
D. meson

4. Which particle has the least mass?
A. electron
B. neutron
C. proton
D. All have the same mass.

5. Cathode rays are composed of
A. alpha particles.
B. electrons.
C. protons.
D. neutrons.

6. The atomic number of an element is the same as the number of
A. protons.
B. neutrons.
C. protons + electrons.
D. protons + neutrons.

7. How many neutrons are present in an atom of tin that has an atomic number of 50 and a mass number of 119?
A. 50
B. 69
C. 119
D. 169

8. What is the mass of 1.50 mol of sodium, Na?
A. 0.652 g
B. 0.478 g
C. 11.0 g
D. 34.5 g

9. How many moles of carbon are in a 28.0 g sample?
A. 336 mol
B. 72.0 mol
C. 2.33 mol
D. 0.500 mol

SHORT ANSWER

10. Which atom has more neutrons, potassium-40 or argon-40?

11. What is the mass of 1.20×10^{23} atoms of phosphorus?

EXTENDED RESPONSE

12. Cathode rays emitted by a piece of silver and a piece of copper illustrate identical properties. What is the significance of this observation?

13. A student believed that she had discovered a new element and named it mythium. Analysis found it contained two isotopes. The composition of the isotopes was 19.9% of atomic mass 10.013 and 80.1% of atomic mass 11.009. What is the average atomic mass, and do you think mythium was a new element?

Test Tip
Choose the best possible answer for each question, even if you think there is another possible answer that is not given.