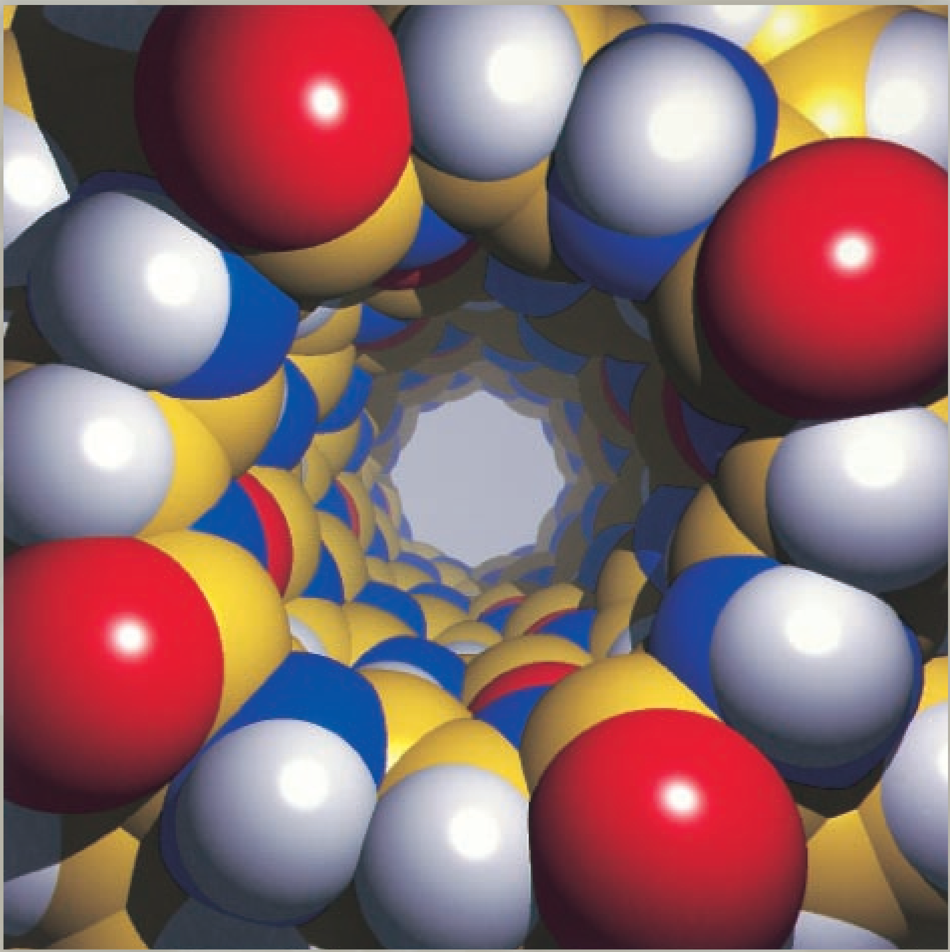
CHAPTER 3

Atoms: The Building Blocks of Matter



An atom is the smallest particle of an element that retains the chemical properties of that element.

The Atom: From Philosophical Idea to Scientific Theory

SECTION 3-1

OBJECTIVES

- Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
- Summarize the five essential points of Dalton's atomic theory.
- Explain the relationship between Dalton's atomic theory and the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.

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 B.C. 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 speculation until the eighteenth century. Then scientists began to gather evidence favoring the atomic theory of matter.

Foundations of Atomic Theory

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. It was also clear that elements combine to form compounds that have different physical and chemical properties than those of the elements that form them. There was great controversy, however, as to whether elements always combine in the same ratio when forming a particular compound.

The transformation of a substance or substances into one or more new substances is known as a *chemical reaction*. 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 lead to the discovery of several basic laws. One of these laws was the **law of conservation of mass,** *which states that mass is neither destroyed nor created 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, *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.**

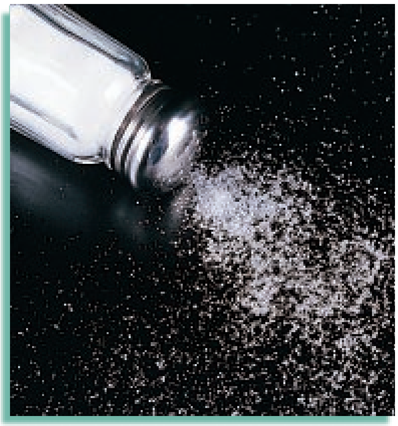


FIGURE 3-1 Each of the salt crystals shown here contains exactly 39.34% sodium and 60.66% chlorine by mass.

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 1.0 g of carbon. In carbon dioxide, 2.66 g of oxygen combine with 1.0 g of carbon. In carbon monoxide, 1.33 g of oxygen combine with 1.0 g of carbon. The ratio of the masses of oxygen in these two compounds is exactly 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.*

Dalton's Atomic Theory

In 1808, an English schoolteacher named John Dalton proposed an explanation for the law of conservation of mass, the law of definite proportions, and the law of multiple proportions. 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 a given 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.

According to Dalton's atomic theory, the law of conservation of mass is explained by the fact that chemical reactions involve merely the combination, separation, or rearrangement of atoms and that during these processes atoms are not subdivided, created, or destroyed. This

internet**connect**

SCI LINKS
NSTA

TOPIC: Atomic theory
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2031

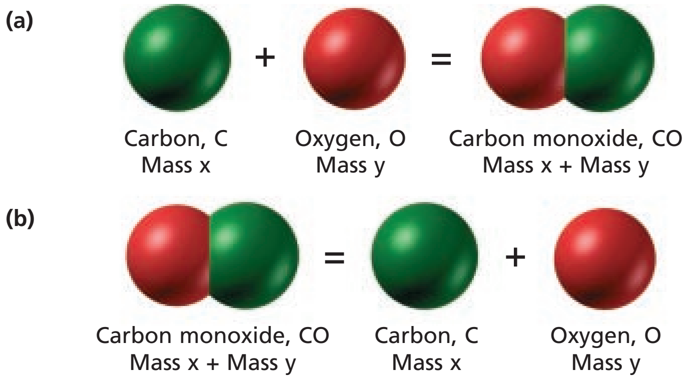


FIGURE 3-2 (a) An atom of carbon, C, and an atom of oxygen, O, can combine chemically to form a molecule of carbon monoxide, CO. The mass of the CO molecule is equal to the mass of the C atom plus the mass of the O atom. (b) The reverse holds true in a reaction in which a CO molecule is broken down into its elements.

idea is illustrated in Figure 3-2 for the formation of carbon monoxide from carbon and oxygen.

The law of definite proportions, on the other hand, results from the fact that a given chemical compound is always composed of the same combination of atoms (see Figure 3-3). 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. This can also be seen in Figure 3-3.

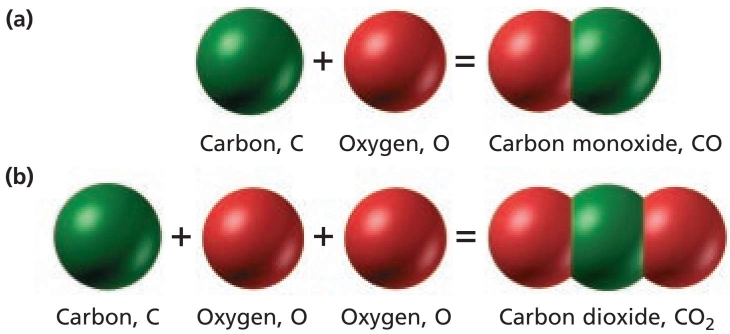


FIGURE 3-3 (a) CO molecules are always composed of one C atom and one O atom. (b) CO_2 molecules are always composed of one C atom and two O atoms. Note that a molecule of carbon dioxide contains twice as many oxygen atoms as does a molecule of carbon monoxide.

Modern Atomic Theory

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 3-3, we know that a given element can have atoms with different masses. Atomic theory has not been discarded, however. Instead, it has been modified to explain the new observations. 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.

internet**connect**

SCI LINKS NSTA

TOPIC: Carbon
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2032

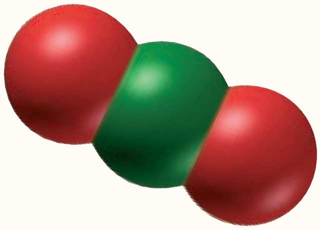
Travels with C

From "Travels with C" by Primo Levi in *Creation to Chaos*

It was to carbon, the element of life, that my first literary dream was turned—and now I want to tell the story of a single atom of carbon.

My fictional character lies, for hundreds of millions of years, bound to three atoms of oxygen and one of calcium, in the form of limestone. (It already has behind it a very long cosmic history, but that we shall ignore.) Time does not exist for it, or exists only in the form of sluggish daily or seasonal variations in temperature. Its existence, whose monotony cannot be conceived of without horror, is an alternation of hots and colds.

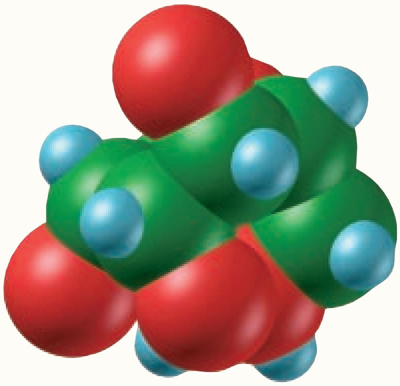
The limestone ledge of which the atom forms a part lies within reach of man and his pickax. At any moment—which I, as narrator, decide out of pure caprice to be the year of 1840—a blow of the pickax detached the limestone and sent it on its way to the lime furnace, where it was plunged into the world of things that change. The atom of carbon was roasted until it separated from the limestone's calcium, which remained, so to speak, with its feet on the ground and went on to meet a less brilliant destiny. Still clinging firmly to two of its three companions, our fictional character issued from the chimney and rode the path of the air. Its story, which once was immobile, now took wing.



Accompanied by two oxygen atoms (red), the carbon atom (green) took to the air.

The atom was caught by the wind, flung down onto the earth, lifted ten kilometers high. It was breathed in by a falcon, but did not penetrate the bird's rich blood and was exhaled. It dissolved three times in the sea, once in the water of a cascading torrent, and again was expelled. It traveled with the wind for eight years—now high, now low, on the sea and among the clouds, over forests, deserts and limitless expanses of ice. Finally, it stumbled into capture and the organic adventure.

The year was 1848. The atom of carbon, accompanied by its two satellites of oxygen, which maintained it in a gaseous state, was borne by the wind along a row of vines. It had the good fortune to brush against a leaf, penetrate it, and be nailed there by a ray of the sun. On entering the leaf, it collided with other innumerable molecules of nitrogen and oxygen. It adhered to a large and complicated molecule that activated it, and simultaneously it received the decisive message from the sky, in the flashing form of a packet of solar light: in an instant, like an insect caught by a spider, the carbon atom was separated from its oxygen, combined with hydrogen, and finally inserted in a chain of life. All this happened swiftly, in silence, at the temperature and pressure of the atmosphere.



Once inside the leaf, the carbon atom joined other carbon atoms (green), as well as hydrogen (blue) and oxygen (red) atoms, to form this molecule essential to life.

Reading for Meaning

Name the various compounds that the carbon atom was a component of during the course of Levi's story.

Read Further

As a component of one particular compound, Levi's carbon atom is breathed in and exhaled by a falcon. Why was it unlikely for the carbon atom to have been taken into the bird's bloodstream?

Quick LAB

Constructing a Model

Wear Safety Goggles and an Apron.

Materials

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

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

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. 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.

SECTION REVIEW

1. Describe the major contributions of each of the following to the modern theory of the atom:
   a. Democritus
   b. John Dalton
2. List the five essential points of Dalton's atomic theory.
3. What chemical laws can be explained on the basis of Dalton's theory?

SECTION 3-2

The Structure of the Atom

OBJECTIVES

- Summarize the observed properties of cathode rays that led to the discovery of the electron.
- Summarize the experiment carried out by Rutherford and his co-workers that led to the discovery of the nucleus.
- List the properties of protons, neutrons, and electrons.
- Define *atom.*

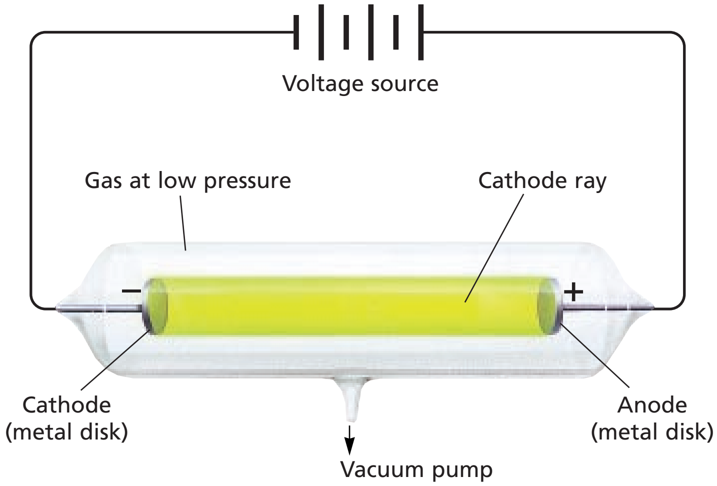
Although John Dalton thought atoms were indivisible, investigators in the late 1800s proved otherwise. As scientific advances allowed a deeper exploration of matter, it became clear that atoms are actually composed of several basic types of smaller particles and that the number and arrangement of these particles within an atom determine that atom's chemical properties. 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 near the center of an atom. In every atom the nucleus contains 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 often referred to as *subatomic particles*.

Discovery of the Electron

The first discovery of a subatomic particle resulted from investigations into the relationship between electricity and matter. In the late 1800s, 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 3-4. Such tubes are known as *cathode-ray tubes*.

FIGURE 3-4 A simple 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 a cathode-ray 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. An object placed between the cathode and the opposite end of the tube cast a shadow on the glass.
2. A paddle wheel placed on rails between the electrodes rolled along the rails from the cathode toward the anode (see Figure 3-5).

These facts supported the existence of a cathode ray. Furthermore, the paddle-wheel experiment showed that a cathode ray had sufficient mass to set the wheel in motion.

Additional experiments provided more information.

3. 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.
4. 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 are composed of identical negatively charged particles, which were later named electrons.

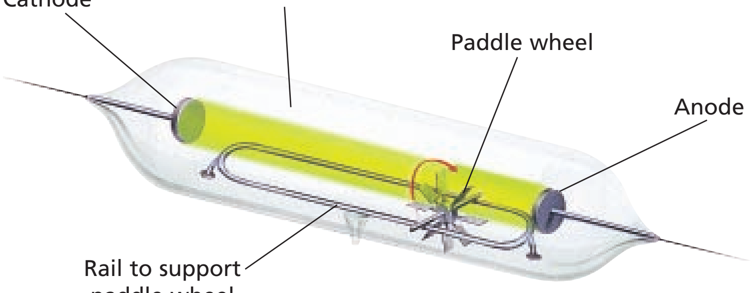


FIGURE 3-5 A paddle wheel placed in the path of the cathode ray moves away from the cathode and toward the anode. The movement of the wheel led scientists to conclude that cathode rays have mass.

Charge and Mass of the Electron

Thomson's experiment revealed that the electron has a very large charge for its tiny mass. In 1909, experiments conducted by the American physicist Robert A. Millikan showed that the mass of the electron is in fact 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.

Millikan's experiments also confirmed that the electron carries a negative electric charge. And because cathode rays have identical properties regardless of the element used to produce them, 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.

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.

CHEMISTRY INTERACTIVE TUTOR
Module 2: Models of the Atom

internet connect
SCLINKS NSTA
TOPIC: Subatomic particles
GO TO: www.scilinks.org
sci LINKS CODE: HC2033

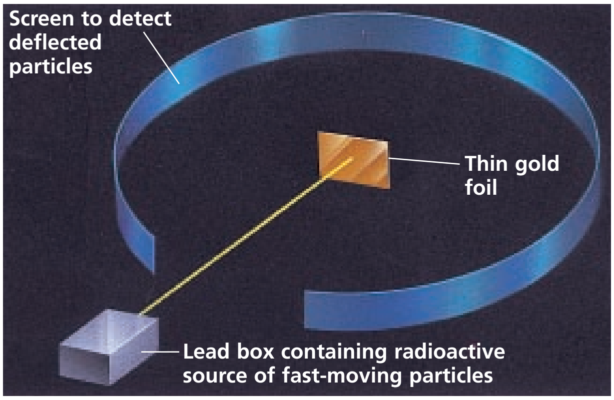
Discovery of the Atomic Nucleus

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, 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. So 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 redirected back toward the source (see Figure 3-6). 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."

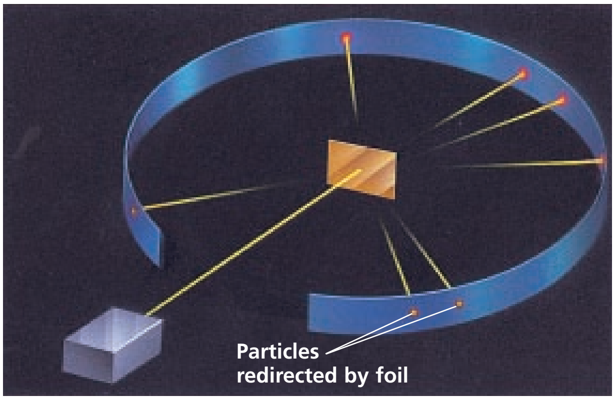
After thinking about the startling result for two years, Rutherford finally came up with an explanation. He reasoned that the rebounded alpha particles must have experienced some powerful force within the atom. And he figured that the source of this force must occupy a very small amount of space because 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 3-7).

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? Although he had no supporting evidence, Rutherford suggested that the electrons surrounded the positively charged nucleus like planets around the sun. He could not explain, however, what kept the electrons in motion around the nucleus.

FIGURE 3-6 (a) Geiger and Marsden bombarded a thin piece of gold foil with a narrow beam of alpha particles. (b) Some of the particles were redirected by the gold foil back toward their source.



(a)



(b)

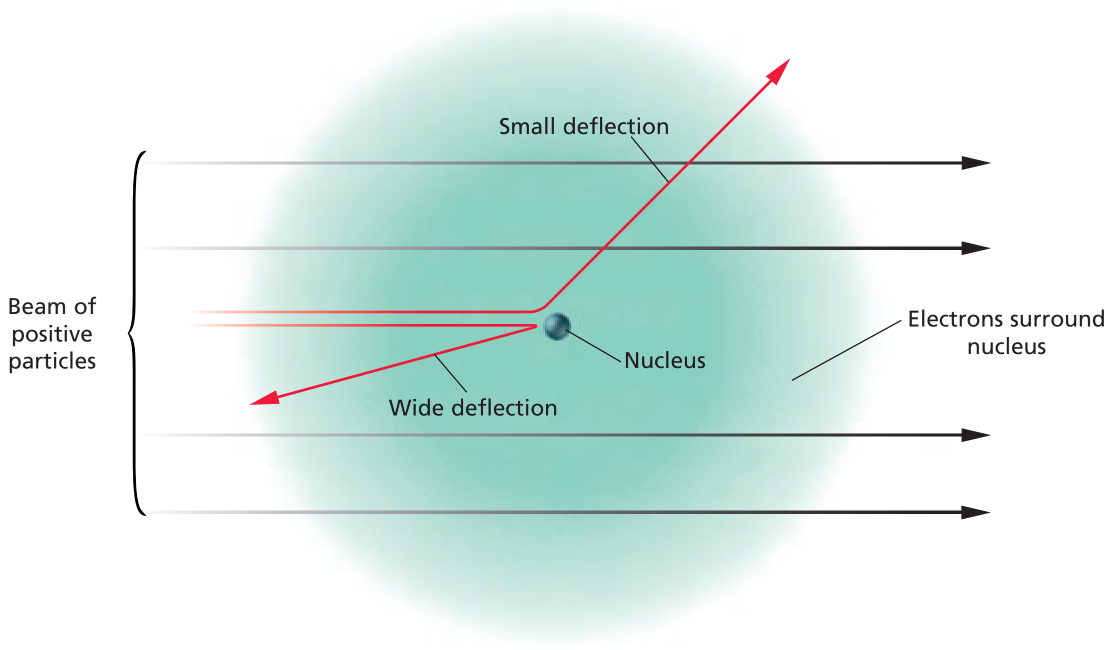


FIGURE 3-7 Rutherford reasoned that each atom in the gold foil contained a small, dense, positively charged nucleus surrounded by electrons. A small number of the alpha particles directed toward the foil were deflected by the tiny nucleus (red arrows). Most of the particles passed through undisturbed (black arrows).

Composition of the Atomic Nucleus

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 contain 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 the number of protons they contain and therefore in the amount of positive charge they possess. Thus, the number of protons in an atom's nucleus 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. Table 3-1 on page 74 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, more than 100

TABLE 3-1 *Properties of Subatomic Particles*

Particle	Symbols	Relative electric charge	Mass number	Relative mass (amu*)	Actual mass (kg)
Electron	e^-, ${}^{\ 0}_{-1}e$	−1	0	0.000 5486	9.109×10^{-31}
Proton	p^+, ${}^1_1\mathrm{H}$	+1	1	1.007 276	1.673×10^{-27}
Neutron	$n°$, ${}^1_0n$	0	1	1.008 665	1.675×10^{-27}

*1 amu (atomic mass unit) = $1.660\ 540 \times 10^{-27}$ kg (see page 78)

protons can exist close together in a 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.**

The Sizes of Atoms

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 REVIEW

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 and contrast the three types of subatomic particles in terms of location in the atom, mass, and relative charge.
4. Why is the cathode-ray tube in Figure 3-4 connected to a vacuum pump?
5. Label the charge on the following in a cathode-ray tube. State the reasons for your answers.
   a. anode
   b. cathode

Counting Atoms

SECTION 3-3

Objectives

- Explain what isotopes are.
- Define *atomic number* and *mass number,* and describe how they apply to isotopes.
- Given the identity of a nuclide, determine its number of protons, neutrons, and electrons.
- Define *mole* in terms of Avogadro's number, and define *molar mass.*
- Solve problems involving mass in grams, amount in moles, and number of atoms of an element.

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.

Atomic Number

All atoms are composed of the same basic 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 in the nucleus of each atom of that element.*

Turn to the inside back cover of this textbook. In the periodic table shown, an element's atomic number is indicated above its symbol. Notice that the elements are placed in order of increasing atomic number. At the top left of the table is hydrogen, H, which has atomic number 1. Atoms of the element hydrogen have one proton in the nucleus. Next in order is helium, He, which has two protons in each nucleus. Lithium, Li, has three protons; beryllium, Be, has four protons; and so on.

The atomic number identifies an element. If you want to know which element has atomic number 47, for example, look at the periodic table. You can see that it is silver, Ag. All silver atoms contain 47 protons in their nuclei. Because atoms are neutral, we know from the atomic number that all silver atoms must also contain 47 electrons.

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

FIGURE 3-8 The atomic number in this periodic-table entry reveals that an atom of lithium has three protons in its nucleus.

Isotopes

The simplest atoms are those of hydrogen. All hydrogen atoms contain only one proton. However, like many naturally occurring elements, hydrogen atoms can contain different numbers of neutrons.

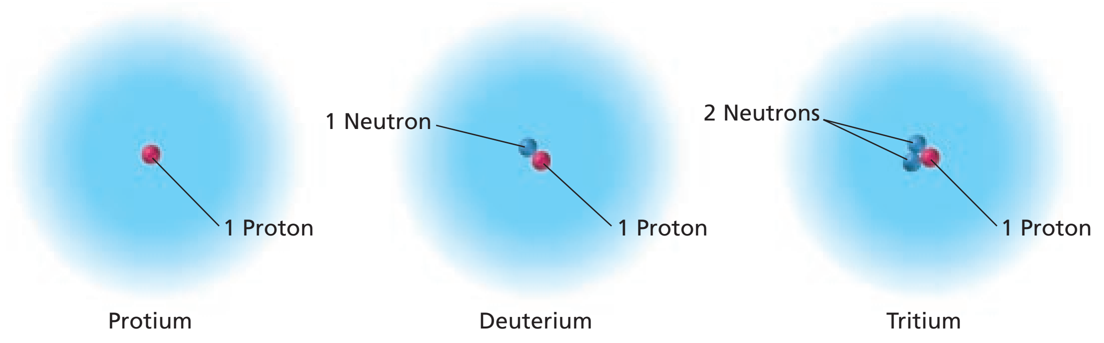


FIGURE 3-9 The nuclei of different isotopes of the same element have the same number of protons but different numbers of neutrons. This is illustrated above by the three isotopes of hydrogen.

Three types of hydrogen atoms are known. The most common type of hydrogen is sometimes called *protium.* It accounts for 99.985% of the hydrogen atoms found on Earth. The nucleus of a protium atom consists of one proton only, and it has one electron moving about it. There are two other known forms of hydrogen. One is called *deuterium,* which accounts for 0.015% of Earth's hydrogen atoms. Each deuterium atom has a nucleus containing one proton and one neutron. The third form of hydrogen is known as *tritium,* which is radioactive. It exists in very small amounts in nature, but it can be prepared artificially. Each tritium atom contains one proton, two neutrons, and one electron.

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, 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.

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 in the nucleus of an isotope.* The three isotopes of hydrogen described earlier have mass numbers 1, 2, and 3, as shown in Table 3-2.

TABLE 3-2 ***Mass Numbers of Hydrogen Isotopes***

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

Designating Isotopes

The isotopes of hydrogen are unusual in that they have distinct names. Isotopes are usually identified by specifying their mass number. There are two methods for specifying isotopes. In the first method, the mass number is written with a hyphen after the name of the element. Tritium, for example, is written as hydrogen-3. We will refer to this method as *hyphen notation*. The uranium isotope used as fuel for nuclear power plants has a mass number of 235 and is therefore known as uranium-235. The second method shows the composition of a nucleus as the isotope's *nuclear symbol*. For example, uranium-235 is written as $^{235}_{92}U$. The superscript indicates the mass number and the subscript indicates the atomic number. The number of neutrons is found by subtracting the atomic number from the mass number.

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

Thus, a uranium-235 nucleus contains 92 protons and 143 neutrons.

Table 3-3 gives the names, symbols, and compositions of the isotopes of hydrogen and helium. **Nuclide** *is a general term for any isotope of any element.* We could say that Table 3-3 lists the compositions of five different nuclides.

internet**connect**

SCILINKS NSTA

TOPIC: Isotopes
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2034

TABLE 3-3 ***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

SAMPLE PROBLEM 3-1

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

SOLUTION

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

3 ***COMPUTE*** The mass number of chlorine-37 is 37. Consulting the periodic table reveals that chlorine's atomic number is 17. The number of neutrons can be found by subtracting the atomic number from the mass number.

$$\text{mass number of chlorine-37} - \text{atomic number of chlorine} = \text{number of neutrons in chlorine-37}$$

$$\text{mass number} - \text{atomic number} = 37 \text{ (protons plus neutrons)} - 17 \text{ protons} = 20 \text{ neutrons}$$

An atom of chlorine-37 contains 17 electrons, 17 protons, and 20 neutrons.

4 ***EVALUATE*** The number of protons in a neutral atom equals the number of electrons. And the sum of the protons and neutrons equals the given mass number.

PRACTICE

1. How many protons, electrons, and neutrons are in an atom of bromine-80?
   Answer
   35 protons, 35 electrons, 45 neutrons
2. Write the nuclear symbol for carbon-13.
   Answer
   $^{13}_{6}C$
3. Write the hyphen notation for the element that contains 15 electrons and 15 neutrons.
   Answer
   phosphorus-30

Relative Atomic Masses

Masses of atoms expressed in grams are very small. As we shall see, an atom of oxygen-16, for example, has a mass of 2.657×10^{-23} g. For most chemical calculations it is more convenient to use *relative* atomic masses. As you read in Chapter 2, 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 relative mass value. The masses of all other atoms are expressed in relation to this defined standard.

The standard used by scientists to govern units of atomic mass is the carbon-12 nuclide. It has been arbitrarily assigned a mass of exactly 12 atomic mass units, or 12 amu. *One* **atomic mass unit,** *or 1 amu, is exactly 1/12 the mass of a carbon-12 atom.* The atomic mass of any nuclide 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 amu. The precise value of the atomic mass of a hydrogen-1 atom is 1.007 825 amu. 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 amu. 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 amu.

Some additional examples of the atomic masses of the naturally occurring isotopes of several elements are given in Table 3-4 on page 80. 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 Table 3-1). The mass of the electron is 0.000 5486 amu, that of the proton is 1.007 276 amu, and that of the neutron is 1.008 665 amu. Note that the proton and neutron masses are close to but not equal to 1 amu. You have learned that the mass number is the total number of protons and neutrons in 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 amu and the atomic masses include electrons. Also, as you will read in Chapter 22, a small amount of mass is changed to energy in the creation of a nucleus from protons and neutrons.

Average Atomic Masses of Elements

Most elements occur naturally as mixtures of isotopes, as indicated in Table 3-4. The percentage of each isotope in the naturally occurring element on Earth is nearly always the same, no matter where the element is found. The percentage at which each of an element's isotopes occurs in nature is taken into account when calculating the element's average atomic mass. **Average atomic mass** *is the weighted average of the atomic masses of the naturally occurring isotopes of an element.*

The following is a simple 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 marbles, 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.

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

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.

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

TABLE 3-4 ***Atomic Masses and Abundances of Several Naturally Occurring Isotopes***

Isotope	Mass number	Percentage natural abundance	Atomic mass (amu)	Average atomic mass of element (amu)
Hydrogen-1	1	99.985	1.007 825	1.007 94
Hydrogen-2	2	0.015	2.014 102	
Carbon-12	12	98.90	12 (by definition)	12.0111
Carbon-13	13	1.10	13.003 355	
Carbon-14	14	trace	14.003 242	
Oxygen-16	16	99.762	15.994 915	15.9994
Oxygen-17	17	0.038	16.999 131	
Oxygen-18	18	0.200	17.999 160	
Copper-63	63	69.17	62.929 599	63.546
Copper-65	65	30.83	64.927 793	
Cesium-133	133	100	132.905 429	132.905
Uranium-234	234	0.005	234.040 947	238.029
Uranium-235	235	0.720	235.043 924	
Uranium-238	238	99.275	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 69.17% copper-63, which has an atomic mass of 62.929 599 amu, and 30.83% copper-65, which has an atomic mass of 64.927 793 amu. 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.

$$0.6917 \times 62.929\,599 \text{ amu} + 0.3083 \times 64.927\,793 \text{ amu} = 63.55 \text{ amu}$$

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

The average atomic mass is included for the elements listed in Table 3-4. 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.*

Relating Mass to Numbers of Atoms

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.

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 order 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 $6.022\ 1367 \times 10^{23}$. This means that exactly 12 g of carbon-12 contains $6.022\ 1367 \times 10^{23}$ 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**—*$6.022\ 1367 \times 10^{23}$—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 (5 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 4 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 figure out 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 amu. The atomic mass of a helium atom is 4.00 amu, 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 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).

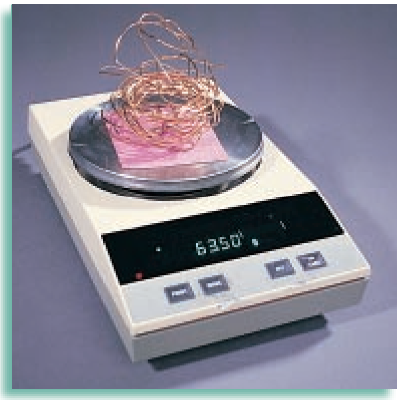
A molar mass of an element contains one mole of atoms. For example, 4.00 g of helium, 6.94 g of lithium, and 200.59 g of mercury all contain a mole of atoms. Figure 3-10 shows molar masses of three common elements.

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

(a)

(b)



(c)

FIGURE 3-10 Shown is approximately one molar mass of each of three elements: (a) carbon (as graphite), (b) iron (nails), and (c) copper (wire).

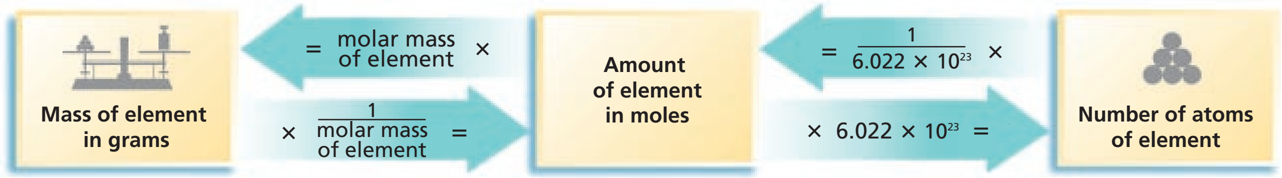


FIGURE 3-11 The diagram shows the relationship between mass in grams, amount in moles, and number of atoms of an element in a sample.

find how many grams of helium there are in two moles of helium, multiply by the molar mass.

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

Figure 3-11 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.

SAMPLE PROBLEM 3-2

What is the mass in grams of 3.50 mol of the element copper, Cu?

SOLUTION

1 ANALYZE

Given: 3.50 mol Cu
Unknown: mass of Cu in grams

2 PLAN

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

According to Figure 3-11, 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 COMPUTE

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

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

4 EVALUATE

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.

PRACTICE

1. What is the mass in grams of 2.25 mol of the element iron, Fe? *Answer* 126 g Fe
2. What is the mass in grams of 0.375 mol of the element potassium, K? *Answer* 14.7 g K
3. What is the mass in grams of 0.0135 mol of the element sodium, Na? *Answer* 0.310 g Na
4. What is the mass in grams of 16.3 mol of the element nickel, Ni? *Answer* 957 g Ni

SAMPLE PROBLEM 3-3

A chemist produced 11.9 g of aluminum, Al. How many moles of aluminum were produced?

SOLUTION

1 ANALYZE

Given: 11.9 g Al
Unknown: amount of Al in moles

2 PLAN

$$\text{mass of Al in grams} \longrightarrow \text{amount of Al in moles}$$

As shown in Figure 3-11, 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 COMPUTE

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

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

4 EVALUATE

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

1. How many moles of calcium, Ca, are in 5.00 g of calcium? *Answer* 0.125 mol Ca
2. How many moles of gold, Au, are in 3.60×10^{-10} g of gold? *Answer* 1.83×10^{-12} mol Au

Conversions with Avogadro's Number

Figure 3-11 shows that 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.

SAMPLE PROBLEM 3-4

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

SOLUTION

1 ANALYZE

Given: 3.01×10^{23} atoms of Ag
Unknown: amount of Ag in moles

2 PLAN

$$\text{number of atoms of Ag} \longrightarrow \text{amount of Ag in moles}$$

From Figure 3-11, 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 Avogrado's number.

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

3 COMPUTE

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

4 EVALUATE

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

PRACTICE

1. How many moles of lead, Pb, are in 1.50×10^{12} atoms of lead?
   Answer 2.49×10^{-12} mol Pb
2. How many moles of tin, Sn, are in 2500 atoms of tin?
   Answer 4.2×10^{-21} mol Sn
3. How many atoms of aluminum, Al, are in 2.75 mol of aluminum?
   Answer 1.66×10^{24} atoms Al

SAMPLE PROBLEM 3-5

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

SOLUTION

1 ANALYZE

Given: 1.20×10^{8} atoms of Cu
Unknown: mass of Cu in grams

2 PLAN

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

As indicated in Figure 3-11, 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 COMPUTE

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

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

4 EVALUATE

Units cancel correctly to give the answer in grams. The size of the answer is reasonable—10^8 has been divided by about 10^{24} and multiplied by about 10^2.

PRACTICE

1. What is the mass in grams of 7.5×10^{15} atoms of nickel, Ni? *Answer* 7.3×10^{-7} g Ni
2. How many atoms of sulfur, S, are in 4.00 g of sulfur? *Answer* 7.51×10^{22} atoms S
3. What mass of gold, Au, contains the same number of atoms as 9.0 g of aluminum, Al? *Answer* 66 g Au

SECTION REVIEW

1. Define 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. $^{64}_{29}Cu$
   d. $^{108}_{47}Ag$
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
   c. 56 electrons and 82 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
7. Determine the number of atoms in the following:
   a. 2.50 mol Zn
   b. 1.50 g C

CHAPTER 3 REVIEW

CHAPTER SUMMARY

3-1

- 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 many chemicals today.
- When elements react to form compounds, they combine in fixed proportions by mass.
- Matter and its mass cannot be created or destroyed in chemical reactions.
- The mass ratio of the elements that make up a given compound is 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.

Vocabulary

law of conservation of mass (66)
law of definite proportions (66)
law of multiple proportions (66)

3-2

- 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—a positively charged, very dense core within the atom—by bombarding metal 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.
- Isotopes of an element differ in the number of neutrons in their nuclei.
- Atomic nuclei have radii of about 0.001 pm (pm = picometers; 1 pm = 10^{-12} m), while atoms have radii of about 40–270 pm.

Vocabulary

atom (70)
nuclear forces (74)

3-3

- The atomic number of an element is equal to the number of protons in the nucleus of an atom of that element.
- The mass number is equal to the total number of protons and neutrons in the nucleus of an atom of that element.
- The relative atomic mass unit (amu) 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 $6.022\,137 \times 10^{23}$. It is equal to the number of atoms in exactly 12 g of carbon-12. A sample that contains a number of particles equal to Avogadro's number contains a mole of those particles.
- The molar mass of an element is the mass of one mole of atoms of that element.

Vocabulary

atomic mass unit (78)
atomic number (75)
average atomic mass (79)
Avogadro's number (81)
isotopes (76)
mass number (76)
molar mass (81)
mole (81)
nuclide (77)

REVIEWING CONCEPTS

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 (3-1)

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? (3-1)

3. a. What is an atom?
b. What two regions make up all atoms? (3-2)

4. Describe at least four properties of electrons that were determined based on the experiments of Thomson and Millikan. (3-2)

5. Summarize Rutherford's model of an atom, and explain how he developed this model based on the results of his famous gold-foil experiment. (3-2)

6. What one number identifies an element? (3-2)

7. a. What are isotopes?
b. How are the isotopes of a particular element alike?
c. How are they different? (3-3)

8. Copy and complete the following table concerning the three isotopes of silicon, Si. (Hint: See Sample Problem 3-1.) (3-3)

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. (3-3)

10. What is a nuclide? (3-3)

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 (3-3)

12. a. What nuclide is used as the standard in the relative scale for atomic masses?
b. What is its assigned atomic mass? (3-3)

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 (3-3)

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? (3-3)

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. (3-3)

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? (3-3)

PROBLEMS

The Mole and Molar Mass

17. What is the mass in grams of each of the following? (Hint: See Sample Problems 3-2 and 3-5.)
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 3-3 and 3-4.)
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

Relative Atomic Mass

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 amu; 0.337%), argon-38 (37.96 amu; 0.063%), and argon-40 (39.96 amu; 99.600%).

20. Naturally occurring boron is 80.20% boron-11 (atomic mass = 11.01 amu) 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 amu average atomic mass of boron? (Write the answer to two decimal places.)

Number of Atoms in a Sample

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

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 an atomic mass unit (amu) related to the mass of a 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 contained in 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. 50.0 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 on pages 70 to 74, 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. Inferring Relationships How much mass is converted into energy during the creation of the nucleus of a $^{235}_{92}U$ nuclide from 92 protons, 143 neutrons, and 92 electrons? (Hint: See Section 22-1.)

TECHNOLOGY & LEARNING

35. **Graphing Calculator** Calculate Numbers of Protons, Electrons, and Neutrons. A graphing calculator can run a program that calculates the numbers of protons, electrons, and neutrons given the atomic mass and numbers for an atom. For example, given a calcium-40 atom, you will calculate the number of protons, electrons, and neutrons in the atom.

Go to Appendix C. If you are using a TI 83 Plus, you can download the program and data sets and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. Remember that you will need to name the program and check the display, as explained in Appendix C. You will then be ready to run the program. After you have graphed the data sets, answer these questions.

a. Which element has the most protons?
b. How many neutrons does mercury-201 have?
c. Carbon-12 and carbon-14 have the same atomic number. Do they have the same number of neutrons? Why or why not?

HANDBOOK SEARCH

36. Group 14 of the *Elements Handbook* 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 instead of CO_2 or O_2?
c. Why is CO poisoning most likely to occur in homes that are well sealed during cold winter months?

RESEARCH & WRITING

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

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

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

40. 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

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

42. **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.