Counting Atoms

SECTION 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, Avogadro's number,* and *molar mass,* and state how all three are related.
- 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 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. 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 8**); 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 the element is silver, Ag. All silver atoms have 47 protons. Because atoms are neutral, we know from the atomic number that all silver atoms must also have 47 electrons.

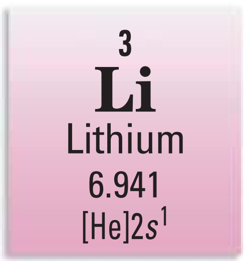

FIGURE 8 The atomic number in this periodic table entry reveals that an atom of lithium has three protons in its nucleus.

Isotopes

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

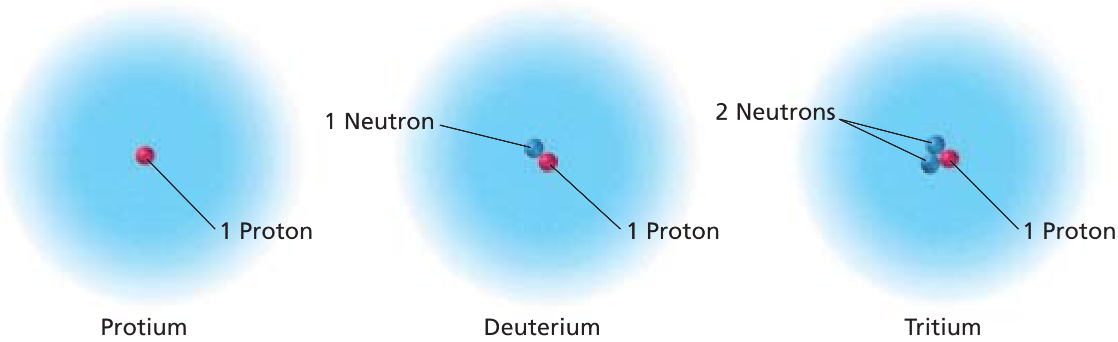

FIGURE 9 The nuclei of different isotopes of the same element have the same number of protons but different numbers of neutrons. This is illustrated above by the three isotopes of hydrogen.

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 with 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 has 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 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 **Table 2.**

TABLE 2 ***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	$1 + 2 = 3$

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 $^{235}_{92}\mathrm{U}$. 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 is made up of 92 protons and 143 neutrons.

Table 3 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.* We could say that **Table 3** lists the compositions of five different nuclides, three hydrogen nuclides and two helium nuclides.

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For a variety of links related to this chapter, go to www.scilinks.org

Topic: Isotopes
SciLinks code: HC60820

TABLE 3 ***Isotopes of Hydrogen and Helium***

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

SAMPLE PROBLEM A

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 is made up of 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 *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.

extension

Go to **go.hrw.com** for more practice problems that ask you to work with numbers of subatomic particles.

Keyword: HC6ATMX

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 compare units of atomic mass is the carbon-12 atom. 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 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 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 4** 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 **Table 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 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 amu and the atomic masses include electrons. Also, as you will read in Chapter 21, a small amount of mass is changed to energy in the creation of a nucleus from its protons and neutrons.

HISTORICAL CHEMISTRY

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, in 1925, a German chemist named Ida Tacke reported the discovery of element 43, which she called masurium, in niobium ores. At the time, her discovery was not accepted because it was thought technetium could not occur naturally. Recent studies confirm that Tacke and coworkers probably did discover element 43.

Average Atomic Masses of Elements

Most elements occur naturally as mixtures of isotopes, as indicated in **Table 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 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.

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

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 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	99.9885	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	
Oxygen-17	17	0.038	16.999 132	15.9994
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 amu, and 30.85% copper-65, which has an atomic mass of 64.927 794 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.

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

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 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 $6.022\,1415 \times 10^{23}$. This means that exactly 12 g of carbon-12 contains $6.022\,1415 \times 10^{23}$ 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**—*$6.022\,1415 \times 10^{23}$—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 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 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).

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 10** shows molar masses of three common elements.


(a)


(b)


(c)

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

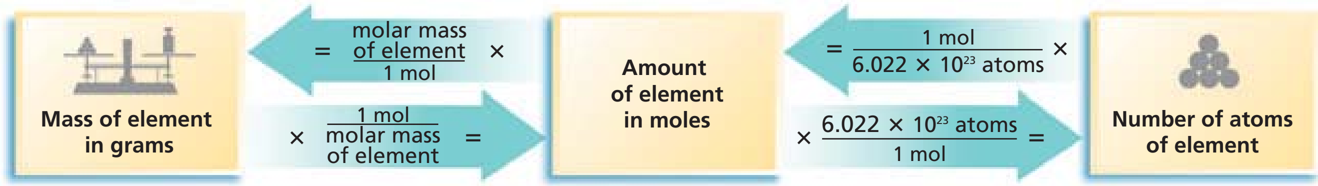

FIGURE 11 The diagram shows the relationship between mass in grams, amount in moles, and number of atoms of an element in a sample.

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.

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

Figure 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 B *For more help, go to the Math Tutor at the end of this chapter.*

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 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.

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

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

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?

extension

Go to **go.hrw.com** for more practice problems that ask you to convert from amount in moles to mass.

Keyword: HC6ATMX

SAMPLE PROBLEM C

For more help, go to the Math Tutor at the end of this chapter.

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 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.

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

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

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?

extension

Go to **go.hrw.com** for more practice problems that ask you to convert from mass to amount in moles.

Keyword: HC6ATMX

Conversions with Avogadro's Number

Figure 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 D *For more help, go to the Math Tutor at the end of this chapter.*

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 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

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

4 EVALUATE The answer is correct—units cancel correctly and the number of atoms is 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?

extension

Go to **go.hrw.com** for more practice problems that ask you to convert between atoms and moles.

Keyword: HC6ATMX

SAMPLE PROBLEM E *For more help, go to the Math Tutor at the end of this chapter.*

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 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.

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

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

Answers in Appendix E

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?

extension

Go to **go.hrw.com** for more practice problems that ask you to convert among atoms, grams, and moles.

Keyword: HC6ATMX

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

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?

CHAPTER HIGHLIGHTS

The Atom: From Philosophical Idea to Scientific Theory

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

- 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.

The Structure of the Atom

Vocabulary
atom
nuclear forces

- 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 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.
- 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.

Counting Atoms

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

- 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 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 $6.022\,1415 \times 10^{23}$. A sample that contains a number of particles equal to Avogadro's number contains a mole of those particles.

CHAPTER REVIEW

For more practice, go to the Problem Bank in Appendix D.

The Atom: From Philosophical Idea to Scientific Theory

SECTION 1 REVIEW

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?

The Structure of the Atom

SECTION 2 REVIEW

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?

Counting Atoms

SECTION 3 REVIEW

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