CHAPTER 5

Counting Atoms and Molecules: The Mole

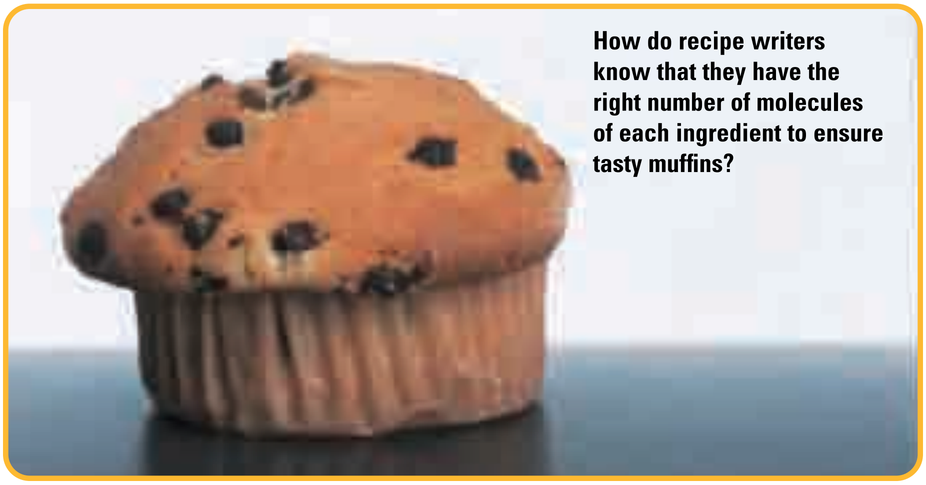
A recipe for chocolate chip muffins tells you exactly what ingredients you will need. One recipe might call for flour, butter, eggs, milk, vinegar, baking soda, sugar, and chocolate chips. It also tells you how much of each ingredient you will need, using convenient units of measurement. Which of the ingredients do you measure by counting? Which do you measure by volume or by mass? The recipe does not tell you exactly how many chocolate chips or grains of sugar you will need. It would take far too long to count individual chocolate chips or grains of sugar. Instead, the amounts are given in millilitres or grams—the units of volume or mass.

In some ways, chemistry is similar to baking. To carry out a reaction successfully—in chemistry or in baking—you need to know how much of each reactant you will need. When you bake something with vinegar and baking soda, for example, the baking soda reacts with acetic acid in the vinegar to produce carbon dioxide gas. The carbon dioxide gas helps the batter rise. The chemical equation for this reaction is

$$NaHCO_{3(s)} + CH_3COOH_{(aq)} \rightarrow NaCH_3COO_{(aq)} + H_2O_{(\ell)} + CO_{2(g)}$$

According to the balanced equation, one molecule of baking soda reacts with one molecule of acetic acid to form a salt, water, and carbon dioxide. If you wanted to carry out the reaction, how would you know the amount of baking soda and vinegar to use? Their molecules are much too small and numerous to be counted like eggs.

In this chapter, you will learn how chemists count atoms by organizing large numbers of them into convenient, measurable groups. You will learn how these groups relate the number of atoms in a substance to its mass. Using your calculator and the periodic table, you will learn how to convert between the mass of a substance and the number of atoms it contains.



How do recipe writers know that they have the right number of molecules of each ingredient to ensure tasty muffins?

Chapter Preview

Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- defining and describing the relationships among atomic number, mass number, atomic mass, and isotope (Chapter 2, section 2.1)
- writing chemical formulas and equations (Chapter 3, section 3.4)
- balancing chemical equations by inspection (Chapter 4, section 4.1)

5.1 Isotopes and Average Atomic Mass

Section Preview/ Specific Expectations

In this section, you will

- **describe** the relationship between isotopic abundance and average atomic mass
- **solve** problems involving percentage abundance of isotopes and relative atomic mass
- **explain** the significance of a weighted average
- **communicate** your understanding of the following terms: *isotopic abundance, average atomic mass, mass spectrometer, weighted average*

How does the mass of a substance relate to the number of atoms in the substance? To answer this question, you need to understand how the relative masses of individual atoms relate to the masses of substances that you can measure on a balance.

The head of a pin, like the one shown in Figure 5.1, is made primarily of iron. It has a mass of about 8×10^{-3} g, yet it contains about 8×10^{19} atoms. Even if you could measure the mass of a single atom on a balance, the mass would be so tiny (about 1×10^{-22} g for an iron atom) that it would be impractical to use in everyday situations. Therefore, you need to consider atoms in bulk, not individually.

How do you relate the mass of individual atoms to the mass of a large, easily measurable number of atoms? In the next two sections, you will find out.

Figure 5.1 The head of a typical pin contains about 80 quintillion atoms.

Relating Atomic Masses to Macroscopic Masses

In Chapter 2, you learned that the mass of an atom is expressed in atomic mass units. Atomic mass units are a relative measure, defined by the mass of carbon-12. According to this definition, one atom of carbon-12 is assigned a mass of 12 u. Stated another way, 1 u = $\frac{1}{12}$ of the mass of one atom of carbon-12.

The masses of all other atoms are defined by their relationship to carbon-12. For example, oxygen-16 has a mass that is 133% of the mass of carbon-12. Hence the mass of an atom of oxygen-16 is $\frac{133}{100} \times 12.000 \text{ u} = 16.0 \text{ u}$.

Usually, not all the atoms in an element have the same mass. As you learned in Chapter 2, atoms of the same element that contain different numbers of neutrons are called isotopes. Most elements are made up of two or more isotopes. Chemists need to account for the presence of isotopes when finding the relationship between the mass of a large number of atoms and the mass of a single atom. To understand why this is important, consider the following analogy.

Imagine that you have the task of finding the total mass of 10 000 spoons. If you know the mass of a dessertspoon, can you assume that its mass represents the average mass of all the spoons? What if the 10 000 spoons include soupspoons, dessertspoons, and tablespoons? If you use the mass of a dessertspoon to calculate the total mass of all the spoons, you may obtain a reasonable estimate. Your answer will not be accurate, however, because each type of spoon has a different mass. You cannot calculate an accurate average mass for all the spoons based on knowing the mass of only one type. How can you improve the accuracy of your answer without determining the mass of all the spoons?

Figure 5.2 **Think about finding the average mass of a group of objects that have different masses. How is this similar to finding the average mass of an element that is composed of different isotopes?**

Isotopic Abundance

Chemists face a situation similar to the one described above. Because all the atoms in a given element do not have the same number of neutrons, they do not all have the same mass. For example, magnesium has three naturally occurring isotopes. It is made up of 79% magnesium-24, 10% magnesium-25, and 11% magnesium-26. Whether the magnesium is found in a supplement tablet (like the ones on the right) or in seawater as $Mg(OH)_2$, it is always made up of these three isotopes in the same proportion. The relative amount in which each isotope is present in an element is called the **isotopic abundance**. It can be expressed as a percent or as a decimal fraction. When chemists consider the mass of a sample containing billions of atoms, they must take the isotopic abundance into account.

CHEM FACT

Magnesium plays a variety of roles in the body. It is involved in energy production, nerve function, and muscle relaxation, to name just a few. The magnesium in these tablets, like all naturally occurring magnesium, is made up of three isotopes.

Average Atomic Mass and the Periodic Table

The **average atomic mass** of an element is the average of the masses of all the element's isotopes. It takes into account the abundance of each isotope within the element. The average atomic mass is the mass that is given for each element in the periodic table.

It is important to interpret averages carefully. For example, in 1996, the average size of a Canadian family was 3.1. Of course, no one family actually has 3.1 people. In the same way, while the average atomic mass of carbon is 12.01 u, no one atom of carbon has a mass of 12.01 u.

Examine Figure 5.3. Since the atomic mass unit is based on carbon-12, why does the periodic table show a value of 12.01 u, instead of exactly 12 u? Carbon is made up of several isotopes, not just carbon-12. Naturally occurring carbon contains carbon-12, carbon-13, and carbon-14. If all these isotopes were present in equal amounts, you could simply find the average of the masses of the isotopes. This average mass would be about 13 u, since the masses of carbon-13 and carbon-14 are about 13 u and 14 u respectively.

The isotopes, though, are not present in equal amounts. Carbon-12 comprises 98.9% of all carbon, while carbon-13 accounts for 1.1%. Carbon-14 is present in a very small amount—about 1×10^{-10}%. It makes sense that the average mass of all the isotopes of carbon is 12.01 u—very close to 12—since carbon-12 is by far the predominant isotope.

CHEM FACT

The only elements with only one naturally occurring isotope are beryllium, sodium, aluminum, and phosphorus.

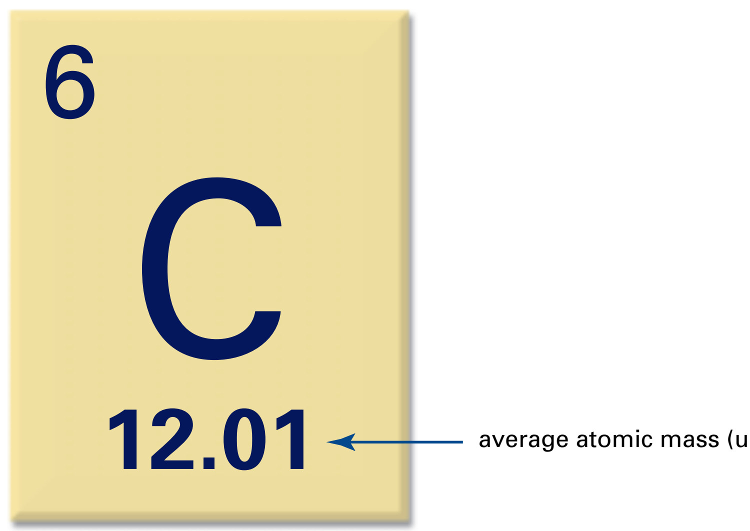


Figure 5.3 **The atomic mass that is given in the periodic table represents the average mass of all the naturally occurring isotopes of the element. It takes into account their isotopic abundances.**

Thus chemists need to know an element's isotopic abundance and the mass of each isotope to calculate the average atomic mass. How do chemists determine the isotopic abundance associated with each element? How do they find the mass of each isotope? They use a **mass spectrometer**, a powerful instrument that generates a magnetic field to obtain data about the mass and abundance of atoms and molecules. You will learn more about the mass spectrometer in Tools & Techniques on page 166. You can use the data obtained with a mass spectrometer to calculate the average atomic mass given in the periodic table.

Working with Weighted Averages

If you obtain the isotopic abundance of an element from mass spectrometer data or a table, you can calculate the average atomic mass of the element. You do this by calculating the **weighted average** of each isotope's mass. A weighted average takes into account not only the values associated with a set of data, but also the abundance or importance of each value.

Normally, when you calculate the average of a set of data, you find the equally weighted average. You add the given values and divide the total by the number of values in the set. Each value in the average is given equal weight. For example, imagine that you have three objects: A, B, and C. A has a mass of 1.0 kg, B has a mass of 2.0 kg, and C has a mass of 3.0 kg. Their average mass is

$$\frac{\text{Mass of }(A + B + C)}{\text{Number of items}} = \frac{1.0\text{ kg} + 2.0\text{ kg} + 3.0\text{ kg}}{3}$$
$$= 2.0\text{ kg}$$

What if you have a set containing two of A, one of B, and three of C? Their average mass becomes

$$\frac{2(1.0\text{ kg}) + 2.0\text{ kg} + 3(3.0\text{ kg})}{6} = 2.2\text{ kg}$$

This is a weighted average.

Another way to calculate the same weighted average is to consider the relative abundance of each object. There are six objects in total. A is present as $\frac{2}{6}$ (33%) of the total, B is present as $\frac{1}{6}$ (17%) of the total, and C is present as $\frac{3}{6}$ (50%) of the total. Thus their average mass can be calculated in the following way:

$$(0.33)(1.0\text{ kg}) + (0.17)(2.0\text{ kg}) + (0.50)(3.0\text{ kg}) = 2.2\text{ kg}$$

mind STRETCH

We use weighted averages all the time! For example, course marks are often based on weighted averages. Suppose that the final mark in a chemistry course is determined as follows: laboratory 25%, tests 30%, homework and quizzes 5%, project 10%, and final exam 30%. A student obtains the following marks: laboratory $\frac{114}{130}$, tests $\frac{261}{300}$, homework and quizzes $\frac{90}{95}$, project $\frac{21}{25}$, and final exam $\frac{70}{80}$. What is the student's final mark in chemistry?

Calculating Average Atomic Mass

You can use a similar method to calculate average atomic mass. If you know the atomic mass of each isotope that makes up an element, as well as the isotopic abundance of each isotope, you can calculate the average atomic mass of the element.

For example, lithium exists as two isotopes: lithium-7 and lithium-6. As you can see in Figure 5.4, lithium-7 has a mass of 7.015 u and makes up 92.58% of lithium. Lithium-6 has a mass of 6.015 u and makes up the remaining 7.42%. To calculate the average atomic mass of lithium, multiply the mass of each isotope by its abundance.

$$\left(\frac{92.58}{100}\right)(7.015\text{ u}) + \left(\frac{7.42}{100}\right)(6.015\text{ u}) = 6.94\text{ u}$$

Looking at the periodic table confirms that the average atomic mass of lithium is 6.94 u. The upcoming Sample Problem gives another example of how to calculate average atomic mass.

CHECKPOINT

How is the atomic mass of an atom different from the mass number of the atom? How are the atomic mass and mass number similar?

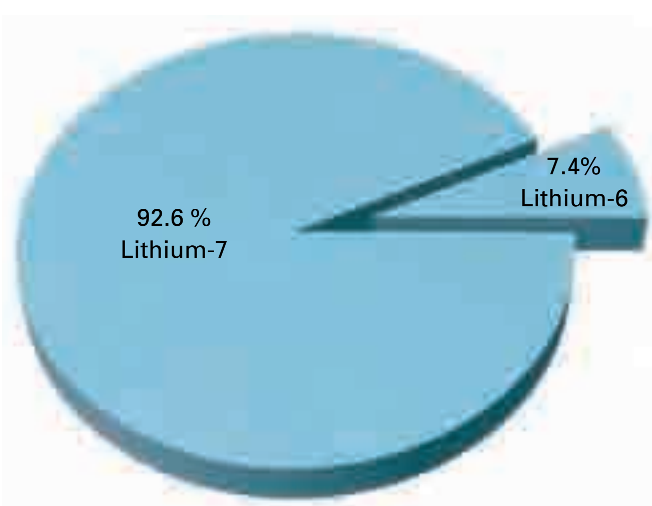


Figure 5.4 Naturally occurring lithium consists of two isotopes, $^{7}_{3}Li$ and $^{6}_{3}Li$.

Tools & Techniques

The Mass Spectrometer

Many chemists depend on instruments known as mass spectrometers. Mass spectrometers can detect trace pollutants in the atmosphere, provide information about the composition of large molecules, and help to determine the age of Earth's oldest rocks.

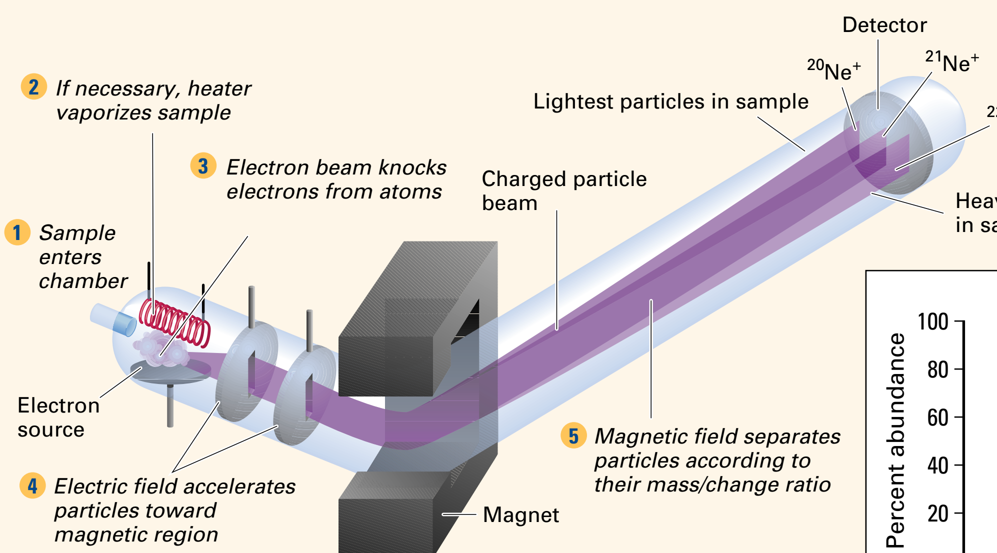
As well, mass spectrometers can find the relative abundance of each isotope in an element. In 1912, J.J. Thompson first detected neon-20 and neon-22 in a sample of neon gas by using a magnetic field to separate the isotopes.

Today's mass spectrometers also use a magnetic field to separate the isotopes of an element. Since a magnetic field can only affect the path of a charged particle, the atoms must first be charged, or "ionized." The magnetic field then deflects ions with the same charge, but different masses, onto separate paths. Imagine rolling a tennis ball at a right angle to the air current from a fan. The air current deflects the path of the ball. The same air current deflects a Ping-Pong™ ball more than a tennis ball. Similarly, a magnetic field deflects isotopes of smaller mass more than isotopes of larger mass.

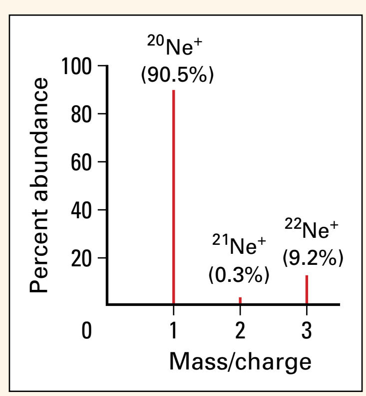
In a mass spectrometer, elements that are not gases are vaporized by heating. Next, the gas atoms are ionized. In electron impact ionization, the gas atoms are bombarded with a stream of electrons from a heated filament. These electrons collide with the gas atoms, causing each atom to lose an electron and become a positive ion.

The ions are focussed and accelerated by electric fields, toward a magnetic field. The magnetic field deflects them and forces them to take a curved path. Lighter ions curve more than heavier ions. Therefore ions from different isotopes arrive at different destinations. Since ions have a charge, a detector can be used to register the current at each destination. The current is proportional to the number of ions that arrive at the destination. An isotope that has a larger relative abundance generates a larger current. From the currents at the different destinations, chemists can deduce the proportion of each isotope in the element.

Mass spectral data for neon, below, show the relative abundance of the isotopes of neon. Using mass spectrometers, chemists now detect a signal from neon-21, in addition to the signals from neon-20 and neon-22.



Mass spectroscopy of neon



Mass spectral data for isotopes of neon

Sample Problem

Average Atomic Mass

Problem

Naturally occurring silver exists as two isotopes. From the mass of each isotope and the isotopic abundance listed below, calculate the average atomic mass of silver.

Isotope	Atomic mass (u)	Relative abundance (%)
$^{107}_{47}Ag$	106.9	51.8
$^{109}_{47}Ag$	108.9	48.2

What Is Required?

You need to find the average atomic mass of silver.

What Is Given?

You are given the relative abundance and the atomic mass of each isotope.

Plan Your Strategy

Multiply the atomic mass of each isotope by its relative abundance, expressed as a decimal. That is, 51.8% expressed as a decimal is 0.518 and 48.2% is 0.482.

Act on Your Strategy

$$\text{Average atomic mass of Ag} = 106.9\text{ u }(0.518) + 108.9\text{ u }(0.482)$$
$$= 107.9\text{ u}$$

Check Your Solution

In this case, the abundance of each isotope is close to 50%. An average atomic mass of about 108 u seems right, because it is between 106.9 u and 108.9 u. Checking the periodic table reveals that the average atomic mass of silver is indeed 107.9 u.

Practice Problems

1. The two stable isotopes of boron exist in the following proportions: 19.78% $^{10}_{5}B$ (10.01 u) and 80.22% $^{11}_{5}B$ (11.01 u). Calculate the average atomic mass of boron.
2. In nature, silicon is composed of three isotopes. These isotopes (with their isotopic abundances and atomic masses) are $^{28}_{14}Si$ (92.23%, 27.98 u), $^{29}_{14}Si$ (4.67%, 28.97 u), and $^{30}_{14}Si$ (3.10%, 29.97 u). Calculate the average atomic mass of silicon.

Continued ...

CHEM FACT

Why are the atomic masses of individual isotopes not exact whole numbers? After all, ^{12}C has a mass of exactly 12 u. Since carbon has 6 neutrons and 6 protons, you might assume that protons and neutrons have masses of exactly 1 u each. In fact, protons and neutrons have masses that are close to, but slightly different from, 1 u. As well, the mass of electrons, while much smaller than the masses of protons and neutrons, must still be taken into account.

CHECKPOINT

Why is carbon-12 the only isotope with an atomic mass that is a whole number?

CHEM FACT

In some periodic tables, the average atomic mass is referred to as the atomic weight of an element. This terminology, while technically incorrect, is still in use and is generally accepted.

Continued ...
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3. Copper is a corrosion-resistant metal that is used extensively in plumbing and wiring. Copper exists as two naturally occurring isotopes: $^{63}_{29}Cu$ (62.93 u) and $^{65}_{29}Cu$ (64.93 u). These isotopes have isotopic abundances of 69.1% and 30.9% respectively. Calculate the average atomic mass of copper.

4. Lead occurs naturally as four isotopes. These isotopes (with their isotopic abundances and atomic masses) are $^{204}_{82}Pb$ (1.37%, 204.0 u), $^{206}_{82}Pb$ (26.26%, 206.0 u), $^{207}_{82}Pb$ (20.82%, 207.0 u), and $^{208}_{82}Pb$ (51.55%, 208.0 u). Calculate the average atomic mass of lead.

ExpressLab A Penny for your Isotopes

The mass of a Canadian penny has decreased several times over the years. Therefore you can use pennies to represent different "isotopes" of a fictitious element, *centium*. That is, each "atom" of *centium* reacts the same way—it is still worth 1¢—but the various isotopes have different characteristic masses.

Safety Precautions

Procedure

1. Obtain a bag of pennies from your teacher. Since the mass of a penny decreased in 1982 and 1997, your bag will contain pennies dated anywhere from 1982 to the present date.
2. Sort your pennies into groups of pre-1997 "isotopes" and post-1997 "isotopes" of *centium*.
3. Count the number of pennies in each group.
4. Find the mass of ten pennies from each group. Divide the total mass by 10 to get the mass of each *centium* "isotope."
5. Use the data you have just gathered to calculate the mass of the pennies, using a weighted average. This represents the "average atomic mass" of *centium*.

Analysis

1. In step 4, you used the average mass of ten pennies to represent the mass of one "isotope" of *centium*.
   (a) Why did you need to do this? Why did you not just find the mass of one penny from each group?
   (b) If you were able to find the mass of real isotopes for this experiment, would you need to do step 4? Explain.
2. Compare your "average atomic mass" for *centium* with the "average atomic mass" obtained by other groups.
   (a) Are all the masses the same? Explain any differences.
   (b) What if you were able to use real isotopes of an element, such as copper, for this experiment? Would you expect results to be consistent throughout the class? Explain.

Calculating Isotopic Abundance

Chemists use a mass spectrometer to determine accurate values for the isotopic abundance associated with each element. Knowing the average atomic mass of an element, you can use the masses of its isotopes to calculate the isotopic abundances.

Sample Problem

Isotopic Abundance

Problem

Boron exists as two naturally occurring isotopes: $^{10}_{5}B$ (10.01 u) and $^{11}_{5}B$ (11.01 u). Calculate the relative abundance of each isotope of boron.

What Is Required?

You need to find the isotopic abundance of boron.

What Is Given?

Atomic mass of $^{10}_{5}B$ = 10.01 u

Atomic mass of $^{11}_{5}B$ = 11.01 u

From the periodic table, the average atomic mass of boron is B = 10.81 u.

Plan Your Strategy

Express the abundance of each isotope as a decimal rather than a percent. The total abundance of both isotopes is therefore 1. Let the abundance of boron-10 be x. Let the abundance of boron-11 be $1 - x$. Set up an equation, and solve for x.

Act on Your Strategy

$$\begin{aligned}
\text{Average atomic mass} &= x(\text{atomic mass B-10}) + (1 - x)(\text{atomic mass B-11}) \\
10.81 &= x(10.01) + (1 - x)(11.01) \\
10.81 &= 10.01x + 11.01 - 11.01x \\
11.01x - 10.01x &= 11.01 - 10.81 \\
x &= 0.2000
\end{aligned}$$

The abundance of boron-10 is 0.2000.

The abundance of boron-11 is $1 - x$, or $1 - 0.2000 = 0.8000$.

The abundance of $^{10}_{5}B$ is therefore 20.00%. The abundance of $^{11}_{5}B$ is 80.00%.

Check Your Solution

The fact that boron-11 comprises 80% of naturally occurring boron makes sense, because the average atomic mass of boron is 10.81 u. This is closer to 11.01 u than to 10.01 u.

Continued ...

CHEM FACT

If you wear contact lenses, you may use boron every day. Boron is part of boric acid, H_3BO_3, which is contained in many cleaning solutions for contact lenses.

Continued ...
FROM PAGE 169

Practice Problems

5. Hydrogen is found primarily as two isotopes in nature: $^{1}_{1}H$ (1.0078 u) and $^{2}_{1}H$ (2.0140 u). Calculate the percentage abundance of each isotope based on hydrogen's average atomic mass.

6. Lanthanum is composed of two isotopes: $^{138}_{57}La$ (137.91 u) and $^{139}_{57}La$ (138.91 u). Look at the periodic table. What can you say about the abundance of $^{138}_{57}La$?

7. Rubidium ignites spontaneously when exposed to oxygen to form rubidium oxide, Rb_2O. Rubidium exists as two isotopes: $^{85}_{37}Rb$ (84.91 u) and $^{87}_{37}Rb$ (86.91 u). If the average atomic mass of rubidium is 85.47 u, determine the percentage abundance of $^{85}_{37}Rb$.

8. Oxygen is composed of three isotopes: $^{16}_{8}O$ (15.995 u), $^{17}_{8}O$ (16.999 u), and $^{18}_{8}O$ (17.999 u). One of these isotopes, $^{17}_{8}O$, comprises 0.037% of oxygen. Calculate the percentage abundance of the other two isotopes, using the average atomic mass of 15.9994 u.

Section Wrap-up

In this section, you learned how isotopic abundance relates to average atomic mass. Since you know the average mass of an atom in any given element, you can now begin to relate the mass of a single atom to the mass of a large number of atoms. First you need to establish how many atoms are in easily measurable samples. In section 5.2, you will learn how chemists group atoms into convenient amounts.

Section Review

1. K/U The average atomic mass of potassium is 39.1 u. Explain why no single atom of potassium has a mass of 39.1 u.

2. I Naturally occurring magnesium exists as a mixture of three isotopes. These isotopes (with their isotopic abundances and atomic masses) are Mg-24 (78.70%, 23.985 u), Mg-25 (10.13%, 24.985 u), and Mg-26 (11.17%, 25.983 u). Calculate the average atomic mass of magnesium.

3. C Assume that an unknown element, X, exists naturally as three different isotopes. The average atomic mass of element X is known, along with the atomic mass of each isotope. Is it possible to calculate the percentage abundance of each isotope? Why or why not?

4. C You know that silver exists as two isotopes: silver-107 and silver-109. However, radioisotopes of silver, such as silver-105, silver-106, silver-108, and silver-110 to silver-117 are known. Why do you not use the abundance and mass of these isotopes when you calculate the average atomic mass of silver? Suggest two reasons.

The Avogadro Constant and the Mole

5.2

In section 5.1, you learned how to use isotopic abundances and isotopic masses to find the average atomic mass of an element. You can use the average atomic mass, found in the periodic table, to describe the average mass of an atom in a large sample.

Why is relating average atomic mass to the mass of large samples important? In a laboratory, as in everyday life, we deal with macroscopic samples. These samples contain incredibly large numbers of atoms or molecules. Can you imagine a cookie recipe calling for six septillion molecules of baking soda? What if copper wire in a hardware store were priced by the atom instead of by the metre, as in Figure 5.5? What if we paid our water bill according to the number of water molecules that we used? The numbers involved would be ridiculously inconvenient. In this section, you will learn how chemists group large numbers of atoms into amounts that are easily measurable.

Section Preview/ Specific Expectations

In this section, you will

- **describe** the relationship between moles and number of particles
- **solve** problems involving number of moles and number of particles
- **explain** why chemists use the mole to group atoms
- **communicate** your understanding of the following terms: *mole, Avogadro constant*

Figure 5.5 Copper wire is often priced by the metre because the metre is a convenient unit. What unit do chemists use to work with large numbers of atoms?

Grouping for Convenience

In a chemistry lab, as well as in other contexts, it is important to be able to measure amounts accurately and conveniently. When you purchase headache tablets from a drugstore, you are confident that each tablet contains the correct amount of the active ingredient. Years of testing and development have determined the optimum amount of the active ingredient that you should ingest. If there is too little of the active ingredient, the tablet may not be effective. If there is too much, the tablet may be harmful. When the tablets are manufactured, the active ingredient needs to be weighed in bulk. When the tablets are tested, however, to ensure that they contain the right amount of the active ingredient, chemists need to know how many molecules of the substance are present. How do chemists group particles so that they know how many are present in a given mass of substance?

On its own, the mass of a chemical is not very useful to a chemist. The chemical reactions that take place depend on the number of atoms present, not on their masses. Since atoms are far too small and numerous to count, you need a way to relate the numbers of atoms to masses that can be measured.

When many items in a large set need to be counted, it is often useful to work with groups of items rather than individual items. When you hear the word "dozen," you think of the number 12. It does not matter what the items are. A dozen refers to the quantity 12 whether the items are eggs or pencils or baseballs. Table 5.1 lists some common quantities that we use to deal with everyday items.

Table 5.1 Some Common Quantities

Item	Quantity	Amount
gloves	pair	2
soft drinks	six-pack	6
eggs	dozen	12
pens	gross (12 dozen)	144
paper	ream	500

You do not buy eggs one at a time. You purchase them in units of a dozen. Similarly, your school does not order photocopy paper by the sheet. The paper is purchased in bundles of 500 sheets, called a ream. It would be impractical to sell sheets of paper individually.

Figure 5.6 **Certain items, because of their size, are often handled in bulk. Would you rather count reams of paper or individual sheets?**

The Definition of the Mole

Convenient, or easily measurable, amounts of elements contain huge numbers of atoms. Therefore chemists use a quantity that is much larger than a dozen or a ream to group atoms and molecules together. This quantity is the **mole** (symbol **mol**).

- One mole (1 mol) of a substance contains $6.022\ 141\ 99 \times 10^{23}$ particles of the substance. This value is called the **Avogadro constant**. Its symbol is N_A.
- The mole is defined as the amount of substance that contains as many elementary entities (atoms, molecules, or formula units) as exactly 12 g of carbon-12.

Language LINK

The term *mole* is an abbreviation of another word. What do you think this other word is? Check your guess by consulting a dictionary.

For example, one mole of carbon contains 6.02×10^{23} atoms of C. One mole of sodium chloride contains 6.02×10^{23} formula units of NaCl. One mole of hydrofluoric acid contains 6.02×10^{23} molecules of HF.

The Avogadro constant is an experimentally determined quantity. Chemists continually devise more accurate methods to determine how many atoms are in exactly 12 g of carbon-12. This means that the accepted value has changed slightly over the years since it was first defined.

Web LINK

www.school.mcgrawhill.ca/resources/
Chemists have devised various ways to determine the Avogadro constant. To learn more about how this constant has been found in the past and how it is found today, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next. What are some methods that chemists have used to determine the number of particles in a mole? How has the accepted value of the Avogadro constant changed over the years?

The Chemist's Dozen

The mole is literally the chemist's dozen. Just as egg farmers and grocers use the dozen (a unit of 12) to count eggs, chemists use the mole (a much larger number) to count atoms, molecules, or formula units. When farmers think of two dozen eggs, they are also thinking of 24 eggs.

$$(2\ \cancel{\text{dozen}}) \times \left(\frac{12\text{ eggs}}{\cancel{\text{dozen}}}\right) = 24\text{ eggs}$$

Chemists work in a similar way. As you have learned above, 1 mol has 6.02×10^{23} particles. Thus 2 mol of aluminum atoms contain 12.0×10^{23} atoms of Al.

$$2\ \cancel{\text{mol}} \times (6.02 \times 10^{23}\ \frac{\text{atoms}}{\cancel{\text{mol}}}) = 1.20 \times 10^{24}\text{ atoms of Al}$$

How Big Is the Avogadro Constant?

The Avogadro constant is a huge number. Its magnitude becomes easier to visualize if you imagine it in terms of ordinary items. For example, suppose that you created a stack of 6.02×10^{23} loonies, as in Figure 5.7. To determine the height of the stack, you could determine the height of one loonie and multiply by 6.02×10^{23}. The Avogadro constant needs to be this huge to group single atoms into convenient amounts. What does 1 mol of a substance look like? Figure 5.8 shows some samples of elements. Each sample contains 6.02×10^{23} atoms. Notice that each sample has a different mass. You will learn why in section 5.3. Examine the following Sample Problem to see how to work with the Avogadro constant.

Figure 5.7 Measure the height of a pile of five loonies. How tall, in kilometres, would a stack of 6.02×10^{23} loonies be?

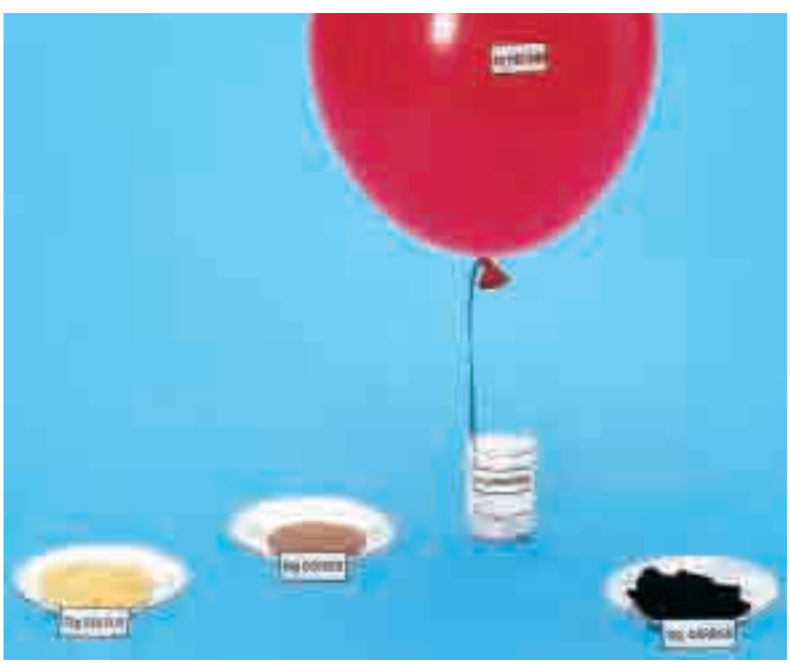

Figure 5.8 Each sample contains 1.00 mol, or 6.02×10^{23} atoms. Why do you think the mass of each sample is different?

Math LINK

Suppose that you invested $\$6.02 \times 10^{23}$ so that it earned 1% compound interest annually. How much money would you have at the end of ten years?

Figure 5.9 **Toronto's SkyDome cost about $500 million to build. Spending $\$6.02 \times 10^{23}$ at the rate of one billion dollars per second is roughly equivalent to building two SkyDomes per second for over 19 million years.**

Sample Problem

Using the Avogadro Constant

Problem

The distance "as the crow flies" from St. John's in Newfoundland to Vancouver in British Columbia is 5046 km. Suppose that you had 1 mol of peas, each of diameter 1 cm. How many round trips could be made between these cities, laying the peas from end to end?

What Is Required?

You need to find the number of round trips from St. John's to Vancouver (2 × 5046 km) that can be made by laying 6.02×10^{23} peas end to end.

What Is Given?

Each round trip is 2 × 5046 km or 10 092 km. A pea has a diameter of 1 cm.

Plan Your Strategy

First convert the round trip distance from kilometres to centimetres. Since each pea has a diameter of 1 cm, a line of 6.02×10^{23} peas is 6.02×10^{23} cm in length. Divide the length of the line of peas by the round-trip distance to find the number of round trips.

Act on Your Strategy

Converting the round-trip distance from kilometres to centimetres gives

$$(10\,092\ \cancel{\text{km}}) \times (10^5\ \text{cm}/\cancel{\text{km}}) = 1.01 \times 10^9\ \text{cm}$$

$$\text{Number of round trips} = \frac{6.02 \times 10^{23}\ \cancel{\text{cm}}}{(1.01 \times 10^9\ \cancel{\text{cm}}/\text{round trip})}$$
$$= 5.96 \times 10^{14}\ \text{round trips}$$

About 596 trillion round trips between St. John's and Vancouver could be made by laying one mole of peas end to end.

Check Your Solution

Looking at the magnitude of the numbers, you have $10^{23} \div 10^9$. This accounts for 10^{14} in the answer.

Practice Problems

9. The length of British Columbia's coastline is 17 856 km. If you laid 6.02×10^{23} metre sticks end to end along the coast of BC, how many rows of metre sticks would you have?

10. The area of Nunavut is 1 936 113 km^2. Suppose that you had 6.02×10^{23} sheets of pastry, each with the dimensions 30 cm × 30 cm. How many times could you cover Nunavut completely with pastry?

Continued ...

Continued ...
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11. If you drove for 6.02×10^{23} days at a speed of 100 km/h, how far would you travel?

12. If you spent \$6.02 × 10^{23} at a rate of \$1.00/s, how long, in years, would the money last? Assume that every year has 365 days.

Converting Moles to Number of Particles

In the Thought Lab below, you can practise working with the mole by relating the Avogadro constant to familiar items. Normally the mole is used to group atoms and compounds. For example, chemists know that 1 mol of barium contains 6.02×10^{23} atoms of Ba. Similarly, 2 mol of barium sulfate contain $2 \times (6.02 \times 10^{23}) = 12.0 \times 10^{23}$ molecules of $BaSO_4$.

number of moles → Multiply by 6.02×10^{23} → **number of particles (atoms, molecules, formula units)**

The mole is used to help us "count" atoms and molecules. The relationship between moles, number of particles, and the Avogadro constant is

$$N = \text{number of particles}$$
$$n = \text{number of moles}$$
$$N_A = \text{Avogadro constant}$$
$$N = n \times N_A$$

Try the next Sample Problem to learn how the number of moles of a substance relates to the number of particles in the substance.

Technology LINK

Scientific calculators are made to accommodate scientific notation easily. To enter the Avogadro constant, for example, type 6.02, followed by the key labelled "EE" or "EXP." (The label on the key depends on the make of calculator you have.) Then enter 23. The 23 will appear at the far right of the display, *without* the exponential base of 10. Your calculator *understands* the number to be in scientific notation.

ThoughtLab The Magnitude of the Avogadro Constant

This activity presents some challenges related to the magnitude of the Avogadro constant. These questions are examples of *Fermi problems*, which involve large numbers (like the Avogadro constant) and give approximate answers. The Italian physicist, Enrico Fermi, liked to pose and solve these types of questions.

Procedure

Work in small groups. Use any reference materials, including materials supplied by your teacher and information on the Internet. For each question, brainstorm to determine the required information. Obtain this information, and answer the question. Be sure to include units throughout your calculation, along with a brief explanation.

Analysis

1. If you covered Canada's land mass with 1.00 mol of golf balls, how deep would the layer of golf balls be?
2. Suppose that you put one mole of five-dollar bills end to end. How many round trips from Earth to the Moon would they make?
3. If you could somehow remove 6.02×10^{23} teaspoons of water from the world's oceans, would you completely drain the oceans? Explain.
4. What is the mass of one mole of apples? How does this compare with the mass of Earth?
5. How many planets would we need for one mole of people, if each planet's population were limited to the current population of Earth?

Language LINK

The term *order of magnitude* refers to the size of a number—specifically to its exponent when in scientific notation. For example, a scientist will say that 50 000 (5×10^4) is two orders of magnitude (10^2 times) larger than 500 (5×10^2). How many orders of magnitude is the Avogadro constant greater than one billion?

Sample Problem

Moles to Atoms

Problem

A sample contains 1.25 mol of nitrogen dioxide, NO_2.

(a) How many molecules are in the sample?

(b) How many atoms are in the sample?

What Is Required?

You need to find the number of atoms and molecules in the sample.

What Is Given?

The sample consists of 1.25 mol of nitrogen dioxide molecules.
Each nitrogen dioxide molecule is made up of three atoms:
1 N atom + 2 O atoms.

$$N_A = 6.02 \times 10^{23} \text{ molecules/mol}$$

Plan Your Strategy

(a) A molecule of NO_2 contains three atoms. Find the number of NO_2 *molecules* in 1.25 mol of nitrogen dioxide.

(b) Multiply the number of molecules by 3 to arrive at the total number of *atoms* in the sample.

Act on Your Strategy

(a) Number of molecules of NO_2

$$= (1.25\ \cancel{\text{mol}}) \times \left(6.02 \times 10^{23}\ \frac{\text{molecules}}{\cancel{\text{mol}}}\right)$$

$$= 7.52 \times 10^{23} \text{ molecules}$$

Therefore there are 7.52×10^{23} molecules in 1.25 mol of NO_2.

(b) $(7.52 \times 10^{23}\ \cancel{\text{molecules}}) \times \left(3\frac{\text{atoms}}{\cancel{\text{molecule}}}\right) = 2.26 \times 10^{24} \text{ atoms}$

Therefore there are 2.26×10^{24} atoms in 1.25 mol of NO_2.

Check Your Solution

Work backwards. One mol contains 6.02×10^{23} atoms. How many moles represent 2.2×10^{24} atoms?

$$2.2 \times 10^{24}\ \cancel{\text{atoms}} \times \frac{1 \text{ mol}}{6.02 \times 10^{23}\ \cancel{\text{atoms}}} = 3.7 \text{ mol}$$

There are 3 atoms in each molecule of NO_2

$$3.7\ \cancel{\text{mol atoms}} \times \frac{1 \text{ mol molecule}}{3\ \cancel{\text{mol atoms}}} = 1.2 \text{ molecules}$$

This is close to the value of 1.25 mol of molecules, given in the question.

Continued ...

Continued ...
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Practice Problems

13. A small pin contains 0.0178 mol of iron, Fe. How many atoms of iron are in the pin?

14. A sample contains 4.70×10^{-4} mol of gold, Au. How many atoms of gold are in the sample?

15. How many formula units are contained in 0.21 mol of magnesium nitrate, $Mg(NO_3)_2$?

16. A litre of water contains 55.6 mol of water. How many molecules of water are in this sample?

17. Ethyl acetate, $C_4H_8O_2$, is frequently used in nail polish remover. A typical bottle of nail polish remover contains about 2.5 mol of ethyl acetate.

(a) How many molecules are in the bottle of nail polish remover?

(b) How many atoms are in the bottle?

(c) How many carbon atoms are in the bottle?

18. Consider a 0.829 mol sample of sodium sulfate, Na_2SO_4.

(a) How many formula units are in the sample?

(b) How many sodium ions, Na^+, are in the sample?

Electronic Learning Partner

Go to the Chemistry 11 Electronic Learning Partner for a video clip that describes the principles behind the Avogadro constant and the mole.

Converting Number of Particles to Moles

Chemists very rarely express the amount of a substance in number of particles. As you have seen, there are far too many particles to work with conveniently. For example, you would never say that you had dissolved 3.21×10^{23} molecules of sodium chloride in water. You might say, however, that you had dissolved 0.533 mol of sodium chloride in water. When chemists communicate with each other about amounts of substances, they usually use units of moles (see Figure 5.10). To convert the number of particles in a substance to the number of moles, rearrange the equation you learned previously.

$$N = N_A \times n$$

$$n = \frac{N}{N_A}$$

To learn how many moles are in a substance when you know how many particles are present, find out how many times the Avogadro constant goes into the number of particles.

Try the next Sample Problem to practise converting the number of atoms, formula units, or molecules in a substance to the number of moles.

Figure 5.10 Chemists rarely use the number of particles to communicate how much of a substance they have. Instead, they use moles.

History LINK

The Avogadro constant is named to honour the Italian chemist Amedeo Avogadro. In 1811, Avogadro postulated what is now known as Avogadro's law: Equal volumes of gases, at equal temperatures and pressures, contain the same number of molecules. You will learn more about Avogadro's law in Unit 4.

Web LINK

www.school.mcgrawhill.ca/resources/
The International System of Units (SI) is based on seven base units for seven quantities. (One of these quantities is the mole.) The quantities are assumed to be all independent of each other. To learn about the seven base quantities, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next. All other SI units are derived from the seven base units.

Sample Problem

Molecules to Moles

Problem

How many moles are present in a sample of carbon dioxide, CO_2, made up of 5.83×10^{24} molecules?

What Is Required?

You need to find the number of moles in 5.83×10^{24} molecules of carbon dioxide.

What Is Given?

You are given the number of molecules in the sample.

$$N_A = 6.02 \times 10^{23} \text{ molecules } CO_2/\text{mol } CO_2$$

Plan Your Strategy

$$n = \frac{N}{N_A}$$

Act on Your Strategy

$$n = \frac{5.83 \times 10^{24} \cancel{\text{ molecules } CO_2}}{(6.02 \times 10^{23} \cancel{\text{ molecules } CO_2}/\text{mol } CO_2)}$$
$$= 9.68 \text{ mol } CO_2$$

There are 9.68 mol of CO_2 in the sample.

Check Your Solution

5.83×10^{24} molecules is approximately equal to 6×10^{24} molecules. Since the number of molecules is about ten times larger than the Avogadro constant, it makes sense that there are about 10 mol in the sample.

Practice Problems

19. A sample of bauxite ore contains 7.71×10^{24} molecules of aluminum oxide, Al_2O_3. How many moles of aluminum oxide are in the sample?

20. A vat of cleaning solution contains 8.03×10^{26} molecules of ammonia, NH_3. How many moles of ammonia are in the vat?

21. A sample of cyanic acid, HCN, contains 3.33×10^{22} atoms. How many moles of cyanic acid are in the sample? **Hint:** Find the number of molecules of HCN first.

22. A sample of pure acetic acid, CH_3COOH, contains 1.40×10^{23} carbon atoms. How many moles of acetic acid are in the sample?

Section Wrap-up

In section 5.1, you learned about the average atomic mass of an element. Then, in section 5.2, you learned how chemists group particles using the mole. In the final section of this chapter, you will learn how to use the average atomic masses of the elements to determine the mass of a mole of any substance. You will learn about a relationship that will allow you to relate the mass of a sample to the number of particles it contains.

Section Review

1. K/U In your own words, define the mole. Use three examples.
2. I Imagine that \$6.02 × 10^{23} were evenly distributed among six billion people. How much money would each person receive?
3. I A typical adult human heart beats an average of 60 times per minute. If you were allotted a mole of heartbeats, how long, in years, could you expect to live? You may assume each year has 365 days.
4. I Calculate the number of atoms in 3.45 mol of iron, Fe.
5. I A sample of carbon dioxide, CO_2, contains 2.56 × 10^{24} molecules.
   - **(a)** How many moles of carbon dioxide are present?
   - **(b)** How many moles of atoms are present?
6. I A balloon is filled with 0.50 mol of helium. How many atoms of helium are in the balloon?
7. I A sample of benzene, C_6H_6, contains 5.69 mol.
   - **(a)** How many molecules are in the sample?
   - **(b)** How many hydrogen atoms are in the sample?
8. I Aluminum oxide, Al_2O_3, forms a thin coating on aluminum when aluminum is exposed to the oxygen in the air. Consider a sample made up of 1.17 mol of aluminum oxide.
   - **(a)** How many molecules are in the sample?
   - **(b)** How many atoms are in the sample?
   - **(c)** How many oxygen atoms are in the sample?
9. C Why do you think chemists chose to define the mole the way they did?
10. I A sample of zinc oxide, ZnO, contains 3.28 × 10^{24} molecules of zinc oxide. A sample of zinc metal contains 2.78 mol of zinc atoms. Which sample contains more zinc: the compound or the element?

Unit Investigation Prep

In your Unit Investigation, you will need to determine the amount of several pure substances in a mixture. Do you think it will be more convenient for you to work with quantities expressed in moles or molecules?

5.3 Molar Mass

Section Preview/ Specific Expectations

In this section, you will

- **explain** the relationship between the average atomic mass of an element and its molar mass
- **solve** problems involving number of moles, number of particles, and mass
- **calculate** the molar mass of a compound
- **communicate** your understanding of the following term: *molar mass*

In section 5.2, you explored the relationship between the number of atoms or particles and the number of moles in a sample. Now you are ready to relate the number of moles to the mass, in grams. Then you will be able to determine the number of atoms, molecules, or formula units in a sample by finding the mass of the sample.

Mass and the Mole

You would never express the mass of a lump of gold, like the one in Figure 5.11, in atomic mass units. You would express its mass in grams. How does the mole relate the number of atoms to measurable quantities of a substance? The definition of the mole pertains to relative atomic mass, as you learned in section 5.1. One atom of carbon-12 has a mass of exactly 12 u. Also, by definition, one mole of carbon-12 atoms (6.02×10^{23} carbon-12 atoms) has a mass of exactly 12 g.

The Avogadro constant is the factor that converts the relative mass of individual atoms or molecules, expressed in atomic mass units, to mole quantities, expressed in grams.

How can you use this relationship to relate mass and moles? The periodic table tells us the average mass of a single atom in atomic mass units (u). For example, zinc has an average atomic mass of 65.39 u. *One mole of an element has a mass expressed in grams numerically equivalent to the element's average atomic mass expressed in atomic mass units*. One mole of zinc atoms has a mass of 65.39 g. This relationship allows chemists to use a balance to count atoms. You can use the periodic table to determine the mass of one mole of an element.

What is Molar Mass?

The mass of one mole of any element, expressed in grams, is numerically equivalent to the average atomic mass of the element, expressed in atomic mass units. The mass of one mole of a substance is called its **molar mass** (symbol $\boldsymbol{M}$). Molar mass is expressed in g/mol. For example, the average atomic mass of gold, as given in the periodic table, is 196.97 u. Thus the mass of one mole of gold atoms, gold's molar mass, is 196.97 g. Table 5.2 gives some additional examples of molar masses.

Table 5.2 Average Atomic Mass and Molar Mass of Four Elements

Element	Average atomic mass (u)	Molar mass (g)
hydrogen, H	1.01	1.01
oxygen, O	16.00	16.00
sodium, Na	22.99	22.99
argon, Ar	39.95	39.95

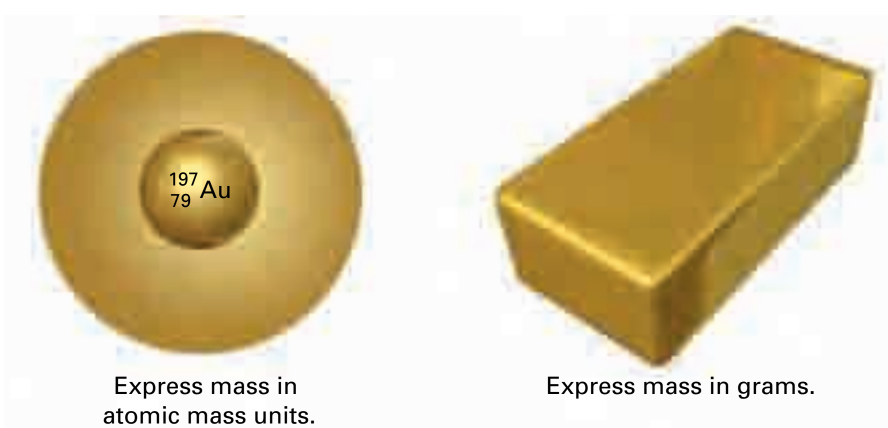

Figure 5.11 **The Avogadro constant is a factor that converts from atomic mass to molar mass.**

Finding the Molar Mass of Compounds

While you can find the molar mass of an element just by looking at the periodic table, you need to do some calculations to find the molar mass of a compound. For example, 1 mol of beryllium oxide, BeO, contains 1 mol of beryllium and 1 mol of oxygen. To find the molar mass of BeO, add the mass of each element that it contains.

$$\begin{aligned} M_{BeO} &= 9.01 \text{ g/mol} + 16.00 \text{ g/mol} \\ &= 25.01 \text{ g/mol} \end{aligned}$$

Examine the following Sample Problem to learn how to determine the molar mass of a compound. Following Investigation 5-A on the next page, there are some Practice Problems for you to try.

CHEM FACT

The National Institute of Standards and Technology (NIST) and most other standardization bodies use *M* to represent molar mass. You may see other symbols, such as *mm*, used to represent molar mass.

Sample Problem

Molar Mass of a Compound

Problem

What is the mass of one mole of calcium phosphate, $Ca_3(PO_4)_2$?

What Is Required?

You need to find the molar mass of calcium phosphate.

What Is Given?

You know the formula of calcium phosphate. You also know, from the periodic table, the average atomic mass of each atom that makes up calcium phosphate.

Plan Your Strategy

Find the total mass of each element to determine the molar mass of calcium phosphate. Find the mass of 3 mol of calcium, the mass of 2 mol of phosphorus, and the mass of 8 mol of oxygen. Then add these masses together.

Act on Your Strategy

$M_{Ca} \times 3 = (40.08 \text{ g/mol}) \times 3 = 120.24 \text{ g/mol}$

$M_P \times 2 = (30.97 \text{ g/mol}) \times 2 = 61.94 \text{ g/mol}$

$M_O \times 8 = (16.00 \text{ g/mol}) \times 8 = 128.00 \text{ g/mol}$

$$\begin{aligned} M_{Ca_3(PO_4)_2} &= 120.24 \text{ g/mol} + 61.94 \text{ g/mol} + 128.00 \text{ g/mol} \\ &= 310.18 \text{ g/mol} \end{aligned}$$

Therefore the molar mass of calcium phosphate is 310.18 g/mol.

Check Your Solution

Using round numbers for a quick check, you get

$$(40 \times 3) + (30 \times 2) + (15 \times 8) = 300$$

This estimate is close to the answer of 310.18 g/mol.

Continued ... ON PAGE 184

PROBLEM TIP

Once you are used to calculating molar masses, you will want to do the four calculations at left all at once. Try solving the Sample Problem using only one line of calculations.

SKILL FOCUS

Modelling concepts

Analyzing and interpreting

Communicating results

Investigation 5-A

Modelling Mole and Mass Relationships

Chemists use the mole to group large numbers of atoms and molecules into manageable, macroscopic quantities. In this way, they can tell how many atoms or molecules are in a given sample, even though the particles are too small to see. In this investigation, you will explore how to apply this idea to everyday objects such as grains of rice and nuts and bolts.

Question

How can you use what you know about the mole and molar mass to count large numbers of tiny objects using mass, and to relate numbers of objects you cannot see based on their masses?

Part 1 Counting Grains of Rice

Materials

electronic balance

40 mL dry rice

50 mL beaker

Procedure

1. Try to measure the mass of a grain of rice. Does the balance register this mass?
2. Count out 20 grains of rice. Measure and record their mass.
3. Find the mass of the empty beaker. Add the rice to the beaker. Find the mass of the beaker and the rice. Determine the mass of the rice.
4. Calculate the number of grains of rice in the 40 mL sample. Report your answer to the number of significant digits that reflects the precision of your calculation.

Part 2 Counting Objects Based on Their Relative Masses

Materials

electronic balance

10 small metal nuts (to represent the fictitious element *nutium*)

10 washers (to represent the fictitious element *washerium*)

2 opaque film canisters with lids

Procedure

1. Measure the mass of 10 nuts (*nutium* atoms). Then measure the mass of 10 washers (*washerium* atoms).
2. Calculate the average mass of a single "atom" of *nutium* and *washerium*.
3. Determine the mass ratio of *nutium* to *washerium*.
4. Obtain, from your teacher, a sealed film canister containing an unknown number of *nutium* atoms. Your teacher will tell you the mass of the empty film canister and lid.
5. Find the mass of the unknown number of *nutium* atoms.
6. You know that you need an equal number of *washerium* atoms to react with the unknown number of *nutium* atoms. What mass of *washerium* atoms do you need?

Analysis

1. Was it possible to get an accurate mass for an individual grain of rice? How did you solve this problem?
2. How did you avoid having to count every single grain of rice in order to determine how many there were in the sample?
3. Using your data, how many grains of rice would be in 6.5×10^3 g of rice?
4. A mole of helium atoms weighs 4.00 g.
   - (a) How many atoms are in 23.8 g of helium?
   - (b) What known relationship did you use to find your answer to part (a)?
   - (c) What analogous relationship did you set up in order to calculate the number of grains of rice based on the mass of the rice?
5. You know the relative mass of nuts and washers. Suppose that you are given some washers in a sealed container. You know that you have the same number of nuts in another sealed container.
   - (a) Can you determine how many washers are in the container without opening either container? Why or why not?
   - (b) What, if any, additional information do you need?
6. The molar mass of carbon is 12.0 g. The molar mass of molecular oxygen is 32.0 g. Equal numbers of carbon atoms and oxygen molecules react to form carbon dioxide.
   - (a) If you have 5.8 g of carbon, what mass of oxygen will react?
   - (b) How does part (a) relate to step 6 in the Procedure for Part 2?

Conclusion

7. How do chemists use the mole and molar masses to count numbers and relative numbers of atoms and molecules? Relate your answer to the techniques you used to count rice, nuts, and washers.

Applications

8. Think about your answer to Analysis question 5(a). Did you need to use the Avogadro constant in your calculation? Explain why or why not.
9. Chemists rarely use the Avogadro constant directly in their calculations. What relationship do they use to avoid working with such a large number?

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Practice Problems

23. State the molar mass of each element.

(a) xenon, Xe

(b) osmium, Os

(c) barium, Ba

(d) tellurium, Te

24. Find the molar mass of each compound.

(a) ammonia, NH_3

(b) glucose, $C_6H_{12}O_6$

(c) potassium dichromate, $K_2Cr_2O_7$

(d) iron(III) sulfate, $Fe_2(SO_4)_3$

25. Strontium may be found in nature as celestite, $SrSO_4$. Find the molar mass of celestite.

26. What is the molar mass of the ion $[Cu(NH_3)_4]^{2+}$?

Counting Particles Using Mass

Using the mole concept and the periodic table, you can determine the mass of one mole of a compound. You know, however, that one mole represents 6.02×10^{23} particles. Therefore you can use a balance to count atoms, molecules, or formula units!

For example, consider carbon dioxide, CO_2. One mole of carbon dioxide has a mass of 44.0 g and contains 6.02×10^{23} molecules. You can set up the following relationship:

$$6.02 \times 10^{23} \text{ molecules of } CO_2 \rightarrow 1 \text{ mol of } CO_2 \rightarrow 44.0 \text{ g of } CO_2$$

How can you use this relationship to find the number of molecules and the number of moles in 22.0 g of carbon dioxide?

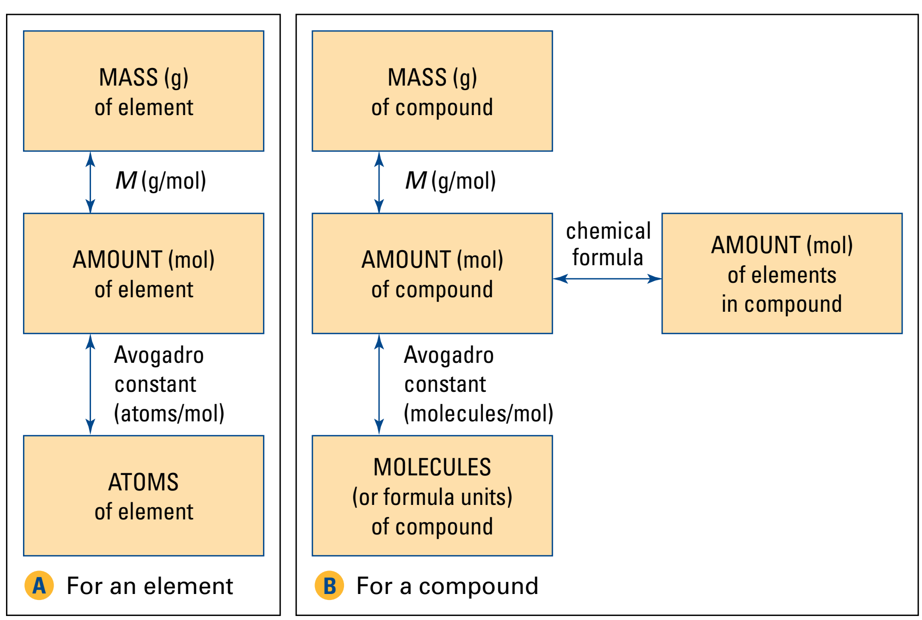


Figure 5.12 The molar mass relates the amount of an element or a compound, in moles, to its mass. Similarly, the Avogadro constant relates the number of particles to the molar amount.

Converting from Moles to Mass

Suppose that you want to carry out a reaction involving ammonium sulfate and calcium chloride. The first step is to obtain one mole of each chemical. How do you decide how much of each chemical you need? You convert the molar amount to mass. Then you use a balance to determine the mass of the proper amount of each chemical.

The following equation can be used to solve problems involving mass, molar mass, and number of moles:

$$\text{Mass} = \text{Number of moles} \times \text{Molar mass}$$
$$m = n \times M$$

$$\boxed{\text{moles (mol)}} \xrightarrow[\text{molar mass (g/mol).}]{\text{Multiply by}} \boxed{\text{mass (g)}}$$

CHECKPOINT

Write the chemical formulas for calcium chloride and ammonium sulfate. Predict what kind of reaction will occur between them. Write a balanced chemical equation to show the reaction. Ionic compounds containing the ammonium ion are soluble. Ammonium sulfate is soluble, but barium chloride is not.

Sample Problem

Moles to Mass

Problem

A flask contains 0.750 mol of carbon dioxide gas, CO_2. What mass of carbon dioxide gas is in this sample?

What Is Required?

You need to find the mass of carbon dioxide.

What Is Given?

The sample contains 0.750 mol. You can determine the molar mass of carbon dioxide from the periodic table.

Plan Your Strategy

In order to convert moles to grams, you need to determine the molar mass of carbon dioxide from the periodic table.

Multiply the molar mass of carbon dioxide by the number of moles of carbon dioxide to determine the mass.

$$m = n \times M$$

Act on Your Strategy

$$M_{CO_2} = 2 \times (16.00 \text{ g/mol}) + 12.01 \text{ g/mol}$$
$$= 44.01 \text{ g/mol}$$

$$m = (0.750 \cancel{\text{mol}}) \times (44.01 \text{ g/}\cancel{\text{mol}})$$
$$= 33.0 \text{ g}$$

The mass of 0.75 mol of carbon dioxide is 33.0 g.

Check Your Solution

1 mol of carbon dioxide has a mass of 44 g. You need to determine the mass of 0.75 mol, or 75% of a mole. 33 g is equal to 75% of 44 g.

Continued ...

Continued ...
FROM PAGE 185

Practice Problems

27. Calculate the mass of each molar quantity.

(a) 3.90 mol of carbon, C

(b) 2.50 mol of ozone, O_3

(c) 1.75×10^7 mol of propanol, C_3H_8O

(d) 1.45×10^{-5} mol of ammonium dichromate, $(NH_4)_2Cr_2O_7$

28. For each group, which sample has the largest mass?

(a) 5.00 mol of C, 1.50 mol of Cl_2, 0.50 mol of $C_6H_{12}O_6$

(b) 7.31 mol of O_2, 5.64 mol of CH_3OH, 12.1 mol of H_2O

29. A litre, 1000 mL, of water contains 55.6 mol. What is the mass of a litre of water?

30. To carry out a particular reaction, a chemical engineer needs 255 mol of styrene, C_8H_8. How many kilograms of styrene does the engineer need?

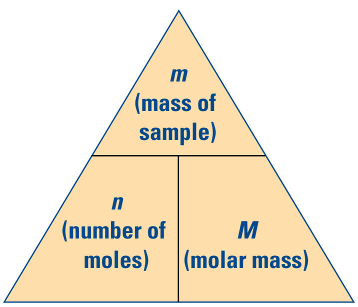


Figure 5.13 Use this triangle for problems involving number of moles, mass of sample, and molar mass. For what other scientific relationships might you use a triangle like this?

You might find the triangle shown in Figure 5.13 useful for problems involving number of moles, number of particles, and molar mass. To use it, cover the quantity that you need to find. The required operation—multiplication or division—will be obvious from the position of the remaining variables. For example, if you want to find the mass of a sample, cover the *m* in the triangle. You can now see that

$$\text{Mass} = \text{Number of moles} \times \text{Molar mass.}$$

Be sure to check that your units cancel.

Converting from Mass to Moles

In the previous Sample Problem, you saw how to convert moles to mass. Often, however, chemists know the mass of a substance but are more interested in knowing the number of moles. Suppose that a reaction produces 223 g of iron and 204 g of aluminum oxide. The masses of the substances do not tell you very much about the reaction. You know, however, that 223 g of iron is 4 mol of iron. You also know that 204 g of aluminum oxide is 2 mol of aluminum oxide. You may conclude that the reaction produces twice as many moles of iron as it does moles of aluminum oxide. You can perform the reaction many times to test your conclusion. If your conclusion is correct, the mole relationship between the products will hold. To calculate the number of moles in a sample, find out how many times the molar mass goes into the mass of the sample.

$$\boxed{\text{moles (mol)}} \xleftarrow[\text{molar mass (g/mol).}]{\text{Divide by}} \boxed{\text{mass (g)}}$$

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

$$n = \frac{m}{M}$$

The following Sample Problem explains how to convert from the mass of a sample to the number of moles it contains.

Sample Problem

Mass to Moles

Problem

How many moles of acetic acid, CH_3COOH, are in a 23.6 g sample?

What Is Required?

You need to find the number of moles in 23.6 g of acetic acid.

What Is Given?

You are given the mass of the sample.

Plan Your Strategy

To obtain the number of moles of acetic acid, divide the mass of acetic acid by its molar mass.

Act on Your Strategy

The molar mass of CH_3COOH is
$(12.01 \times 2) + (16.00 \times 2) + (1.01 \times 4) = 60.06$ g.

$$n = \frac{m}{M_{CH_3COOH}}$$

$$n \text{ mol } CH_3COOH = \frac{23.6 \cancel{\text{g}}}{60.06 \cancel{\text{g}}/\text{mol}}$$

$$= 0.393 \text{ mol}$$

Therefore there are 0.393 mol of acetic acid in 23.6 g of acetic acid.

Check Your Solution

Work backwards. There are 60.06 g in each mol of acetic acid. So in 0.393 mol of acetic acid, you have $0.393 \cancel{\text{mol}} \times 60.06 \text{ g}/\cancel{\text{mol}} = 23.6$ g of acetic acid. This value matches the question.

Practice Problems

31. Calculate the number of moles in each sample.

(a) 103 g of Mo

(b) 1.32×10^4 g of Pd

(c) 0.736 kg of Cr

(d) 56.3 mg of Ge

32. How many moles of compound are in each sample?

(a) 39.2 g of silicon dioxide, SiO_2

(b) 7.34 g of nitrous acid, HNO_2

(c) 1.55×10^5 kg of carbon tetrafluoride, CF_4

(d) 8.11×10^{-3} mg of 1-iodo-2,3-dimethylbenzene C_8H_9I

33. Sodium chloride, NaCl, can be used to melt snow. How many moles of sodium chloride are in a 10 kg bag?

34. Octane, C_8H_{18}, is a principal ingredient of gasoline. Calculate the number of moles in a 20.0 kg sample of octane.

Chemistry Bulletin

Science | Technology | Society | Environment

Chemical Amounts in Vitamin Supplements

Vitamins and minerals (micronutrients) help to regulate your metabolism. They are the building blocks of blood and bone, and they maintain muscles and nerves. In Canada, a standard called *Recommended Nutrient Intake* (*RNI*) outlines the amounts of micronutrients that people should ingest each day. Eating a balanced diet is the best way to achieve your RNI. Sometimes, however, you may need to take multivitamin supplements when you are unable to attain your RNI through diet alone.

The label on a bottle of supplements lists all the vitamins and minerals the supplements contain. It also lists the form and source of each vitamin and mineral, and the amount of each. The form of a mineral is especially important to know because it affects the quantity your body can use. For example, a supplement may claim to contain 650 mg of calcium carbonate, $CaCO_3$, per tablet. This does not mean that there is 650 mg of calcium. The amount of actual calcium, or *elemental calcium*, in calcium carbonate is only 260 mg. Calcium carbonate has more elemental calcium than the same amount of calcium gluconate, which only has 58 mg for every 650 mg of the compound. Calcium gluconate may be easier for your body to absorb, however.

Quality Control

Multivitamin manufacturers employ chemists, or analysts, to ensure that the products they make have the right balance of micronutrients. Manufacturers have departments devoted to *quality control* (*QC*). QC chemists analyze all the raw materials in the supplements, using standardized tests. Most manufacturers use tests approved by a "standardization body," such as the US Pharmacopoeia. Such standardization bodies have developed testing guidelines to help manufacturers ensure that their products contain what the labels claim, within strict limits.

To test for quality, QC chemists prepare samples of the raw materials from which they will make the supplements. They label the samples according to the "lot" of materials from which the samples were taken. They powder and weigh the samples. Then they extract the vitamins. At the same time, they prepare standard solutions containing a known amount of each vitamin.

Next the chemists compare the samples to the standards by subjecting both to the same tests. One test that is used is *high-performance liquid chromatography* (*HPLC*). HPLC produces a spectrum, or "fingerprint," that identifies each compound. Analysts compare the spectrum that is produced by the samples to the spectrum that is produced by the standard.

Analysts test tablets and capsules for dissolution and disintegration properties. The analyst may use solutions that simulate the contents of the human stomach or intestines for these tests. Only when the analysts are sure that the tablets pass all the necessary requirements are the tablets shipped to retail stores.

Making Connections

1. Why might consuming more of the daily RNI of a vitamin or mineral be harmful?
2. The daily RNI of calcium for adolescent females is 700 to 1100 mg. A supplement tablet contains 950 mg of calcium citrate. Each gram of calcium citrate contains 5.26×10^{-3} mol calcium. How many tablets would a 16-year-old female have to take to meet her daily RNI?

An analyst tests whether tablets will sufficiently dissolve within a given time limit.

Converting Between Moles, Mass, and Number of Particles

You can use what you now know about the mole to carry out calculations involving molar mass and the Avogadro constant. One mole of any compound or element contains 6.02×10^{23} particles. The compound or element has a mass, in grams, that is determined from the periodic table.

Now that you have learned how the number of particles, number of moles, and mass of a substance are related, you can convert from one value to another. Usually chemists convert from moles to mass and from mass to moles. Mass is a property that can be measured easily. The following graphic shows the factors used to convert between particles, moles, and mass. Moles are a convenient way to communicate the amount of a substance.

$$\text{number of particles} \xleftrightarrow[\text{constant}]{\text{Avogadro}} \text{moles (mol)} \xleftrightarrow[\text{(g/mol)}]{\text{molar mass}} \text{mass (g)}$$

For example, suppose that you need 2.3 mol of potassium chloride to carry out a reaction. You need to convert the molar amount to mass so that you can measure the correct amount with a balance.

To be certain you understand the relationship among particles, moles, and mass, examine the following Sample Problem.

Sample Problem

Particles to Mass

Problem

What is the mass of 5.67×10^{24} molecules of cobalt(II) chloride, $CoCl_2$?

What Is Required?

You need to find the mass of 5.67×10^{24} molecules of cobalt(II) chloride.

What Is Given?

You are given the number of molecules.

Plan Your Strategy

Convert the number of molecules into moles by dividing by the Avogadro constant. Then convert the number of moles into grams by multiplying by the molar mass of cobalt(II) chloride.

Act on Your Strategy

$$\frac{\text{Number of molecules } CoCl_2}{\text{Number of molecules } CoCl_2/\text{mol } CoCl_2} \times \text{mass } CoCl_2/\text{mol } CoCl_2$$

$$= \frac{5.67 \times 10^{24}\ \cancel{\text{molecules } CoCl_2}}{6.02 \times 10^{23}\ \cancel{\text{molecules } CoCl_2}/\cancel{\text{mol } CoCl_2}} \times 129.84\ \text{g } CoCl_2/\cancel{\text{mol } CoCl_2}$$

$$= 1.22 \times 10^{3}\ \text{g } CoCl_2$$

Continued ...

Math LINK

Average atomic mass values in some periodic tables can have five or more significant digits. How do you know how many significant digits to use? When you enter values, such as average atomic mass, into your calculator, be sure that you use at least one more significant digit than is required in your answer.

Continued ...
FROM PAGE 189

Check Your Solution

5.67×10^{24} molecules is roughly 10 times the Avogadro constant. This means that you have about 10 mol of cobalt(II) chloride. The molar mass of cobalt(II) chloride is about 130 g, and 10 times 130 g is 1300 g.

Practice Problems

35. Determine the mass of each sample.

(a) 6.02×10^{24} formula units of $ZnCl_2$

(b) 7.38×10^{21} formula units of $Pb_3(PO_4)_2$

(c) 9.11×10^{23} molecules of $C_{15}H_{21}N_3O_{15}$

(d) 1.20×10^{29} molecules of N_2O_5

36. What is the mass of lithium in 254 formula units of lithium chloride, LiCl?

37. Express the mass of a single atom of titanium, Ti, in grams.

38. Vitamin B_2, $C_{17}H_{20}N_4O_6$, is also called riboflavin. What is the mass, in grams, of a single molecule of riboflavin?

What if you wanted to compare amounts of substances, and you only knew their masses? You would probably convert their masses to moles. The Avogadro constant relates the molar amount to the number of particles. Examine the next Sample Problem to learn how to convert mass to number of particles.

Sample Problem

Mass to Particles

Problem

Chlorine gas, Cl_2, can react with iodine, I_2, to form iodine chloride, ICl. How many molecules of iodine chloride are contained in a 2.74×10^{-1} g sample?

What Is Required?

You need to find the number of molecules in 2.74×10^{-1} g of iodine.

What Is Given?

You are given the mass of the sample.

Continued ...

Continued ...
FROM PAGE 190

Plan Your Strategy

First convert the mass to moles, using the molar mass of iodine. Multiplying the number of moles by the Avogadro constant will yield the number of molecules.

Act on Your Strategy

The molar mass of ICl is 162.36 g.
Dividing the given mass of ICl by the molar mass gives

$$n = \frac{2.74 \times 10^{-1}\ \cancel{g}}{162.36\ \cancel{g}/\text{mol}}$$
$$= 1.69 \times 10^{-3}\ \text{mol}$$

Now multiply the number of moles by the Avogadro constant. This gives the number of molecules in the sample.

$$(1.69 \times 10^{-3}\ \cancel{\text{mol}}) \times \frac{(6.02 \times 10^{23}\ \text{molecules})}{1\ \cancel{\text{mol}}} = 1.01 \times 10^{21}\ \text{molecules}$$

Therefore there are 1.01×10^{21} molecules in 2.74×10^{-1} g of iodine chloride.

Check Your Solution

Work backwards. Each mole of iodine chloride has a mass of 162.36 g/mol. Therefore 1.01×10^{21} molecules of iodine chloride have a mass of:

$$1.01 \times 10^{21}\ \cancel{\text{molecules}} \times \frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}} \times \frac{162.36\ \text{g}}{1\ \cancel{\text{mol}}}$$
$$= 2.72 \times 10^{-1}\ \text{g}$$

The answer is close to the value given in the question. Your answer is reasonable.

Practice Problems

39. Determine the number of molecules or formula units in each sample.

(a) 10.0 g of water, H_2O

(b) 52.4 g of methanol, CH_3OH

(c) 23.5 g of disulfur dichloride, S_2Cl_2

(d) 0.337 g of lead(II) phosphate, $Pb_3(PO_4)_2$

40. How many atoms of hydrogen are in 5.3×10^4 molecules of sodium glutamate, $NaC_5H_8NO_4$?

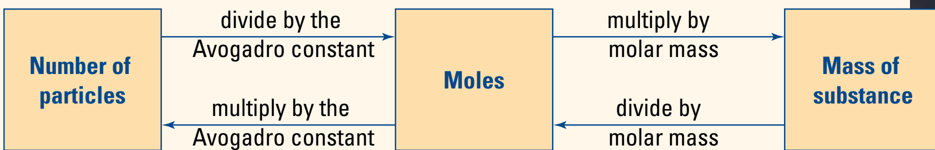
41. How many molecules are in a 64.3 mg sample of tetraphosphorus decoxide, P_4O_{10}?

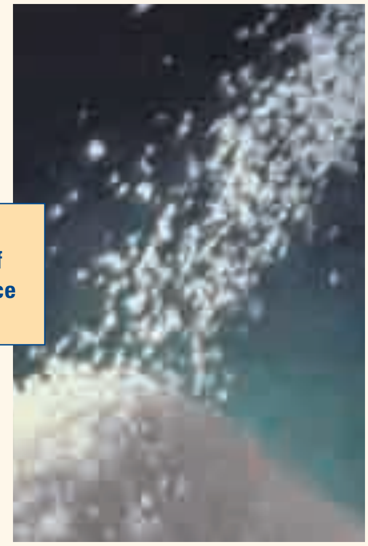
42. (a) How many formula units are in a 4.35×10^{-2} g sample of potassium chlorate, $KClO_3$?

(b) How many ions (chlorate and potassium) are in this sample?

Concept Organizer

Relationships Among Particles, Moles, and Mass





Section Wrap-up

In this chapter, you have learned about the relationships among the number of particles in a substance, the amount of a substance in moles, and the mass of a substance. Given the mass of any substance, you can now determine how many moles and particles make it up. In the next chapter, you will explore the mole concept further. You will learn how the mass proportions of elements in compounds relate to their formulas.

Section Review

Unit Investigation Prep

Before you design your experiment to determine the composition of a mixture, be sure you understand the relationship between moles and mass.

1. C Draw a diagram that shows the relationship between the atomic mass and molar mass of an element and the Avogadro constant.

2. I Consider a 78.6 g sample of ammonia, NH_3.
   (a) How many moles of ammonia are in the sample?
   (b) How many molecules of ammonia are in the sample?

3. I Use your understanding of the mole to answer the following questions.
   (a) What is the average mass, in grams, of a single atom of silicon, Si?
   (b) What is the mass, in atomic mass units, of a mole of silicon atoms?

4. I Consider a 0.789 mol sample of sodium chloride, NaCl.
   (a) What is the mass of the sample?
   (b) How many formula units of sodium chloride are in the sample?
   (c) How many ions are in the sample?

5. I A 5.00 carat diamond has a mass of 1.00 g. How many carbon atoms are in a 5.00 carat diamond?

6. I A bottle of mineral supplement tablets contains 100 tablets and 200 mg of copper. The copper is found in the form of cupric oxide. What mass of cupric oxide is contained in each tablet?

CHAPTER 5 Review

Reflecting on Chapter 5

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Describe the relationships among isotopic abundance, isotopic masses, and average atomic mass.
- Explain why you need to use a weighted average to calculate average atomic mass.
- Calculate isotopic abundance based on isotopic masses and average atomic mass.
- Explain how chemists use a mass spectrometer to determine isotopic abundance and the masses of different isotopes.
- Describe how and why chemists group atoms and molecules into molar amounts.
- Explain how chemists define the mole and why this definition is useful.
- Use the Avogadro constant to convert between moles and particles.
- Explain the relationship between average atomic mass and the mole.
- Find a compound's molar mass using the periodic table.
- Convert between particles, moles, and mass.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

average atomic mass	Avogadro constant
isotopic abundance	mass spectrometer
mole	molar mass
weighted average	

Knowledge/Understanding

1. Distinguish between atomic mass and average atomic mass. Give examples to illustrate each term.
2. Explain why you use a weighted average, based on the masses and abundances of the isotopes, to calculate the average atomic mass of an element.
3. The periodic table lists the average atomic mass of chlorine as 35.45 u. Are there any chlorine atoms that have a mass of 35.45 u? Explain your answer.
4. Explain how the Avogadro constant, average atomic mass, and molar mass are related.
5. Explain how a balance allows a chemist to count atoms or molecules indirectly.
6. (a) Describe the relationship between the mole, the Avogadro constant, and carbon-12.
   (b) Why do chemists use the concept of the mole to deal with atoms and molecules?
7. How is the molar mass of an element related to average atomic mass?
8. Explain what the term molar mass means for each of the following, using examples.
   (a) a metallic element
   (b) a diatomic element
   (c) a compound

Inquiry

9. The isotopes of argon have the following relative abundances: Ar-36, 0.34%; Ar-38, 0.06%; and Ar-40, 99.66%. Estimate the average atomic mass of argon.
10. The isotopes of gallium have the following relative abundances: Ga-69, 60.0%; and Ga-71, 40.0%. Estimate the average atomic mass of gallium using the mass numbers of the isotopes.
11. Estimate the average atomic mass of germanium, given its isotopes with their relative abundances: Ge-70 (20.5%), Ge-72 (27.4%), Ge-73 (7.8%), Ge-74 (36.5%), and Ge-76 (7.8%).
12. Potassium exists as two naturally occurring isotopes: K-39 and K-41. These isotopes have atomic masses of 39.0 u and 41.0 u respectively. If the average atomic mass of potassium is 39.10 u, calculate the relative abundance of each isotope.
13. How many moles of the given substance are present in each sample below?
    (a) 0.453 g of Fe_2O_3
    (b) 50.7 g of H_2SO_4
    (c) 1.24×10^{-2} g of Cr_2O_3
    (d) 8.2×10^{2} g of $C_2Cl_3F_3$
    (e) 12.3 g of NH_4Br

14. Convert each quantity to an amount in moles.
 (a) 4.27×10^{21} atoms of He
 (b) 7.39×10^{23} molecules of ICl
 (c) 5.38×10^{22} molecules of NO_2
 (d) 2.91×10^{23} formula units of $Ba(OH)_2$
 (e) 1.62×10^{24} formula units of KI
 (f) 5.58×10^{20} molecules of C_3H_8

15. Copy the following table into your notebook and complete it.

Sample	Molar mass (g/mol)	Mass of sample (g)	Formula units or molecules	Amount of molecules (mol)	Amount of atoms (mol)
NaCl	58.4	58.4	6.02×10^{23}	1.00	2.00
NH_3		24.8			
H_2O			5.28×10^{22}		
Mn_2O_3					0.332
K_2CrO_4		9.67×10^{-1}			
$C_8H_8O_3$			7.90×10^{24}		
$Al(OH)_3$				8.54×10^{2}	

16. Calculate the molar mass of each compound.
 (a) $PtBr_2$
 (b) $C_3H_5O_2H$
 (c) Na_2SO_4
 (d) $(NH_4)_2Cr_2O_7$
 (e) $Ca_3(PO_4)_2$
 (f) Cl_2O_7

17. Express each quantity as a mass in grams.
 (a) 3.70 mol of H_2O
 (b) 8.43×10^{23} molecules of PbO_2
 (c) 14.8 mol of $BaCrO_4$
 (d) 1.23×10^{22} molecules of Cl_2
 (e) 9.48×10^{23} molecules of HCl
 (f) 7.74×10^{19} molecules of Fe_2O_3

18. How many atoms of C are contained in 45.6 g of C_6H_6?

19. How many atoms of F are contained in 0.72 mol of BF_3?

20. Calculate the following.
 (a) the mass (u) of one atom of xenon
 (b) the mass (g) of one mole of xenon atoms
 (c) the mass (g) of one atom of xenon
 (d) the mass (u) of one mole of xenon atoms
 (e) the number of atomic mass units in one gram

21. How many atoms of C are in a mixture containing 0.237 mol of CO_2 and 2.38 mol of CaC_2?

22. How many atoms of H are in a mixture of 3.49×10^{23} molecules of H_2O and 78.1 g of CH_3OH?

23. How many nitrate ions are in a solution that contains 3.76×10^{-1} mol of calcium nitrate, $Ca(NO_3)_2$?

24. Ethanol, C_2H_5OH, is frequently used as the fuel in wick-type alcohol lamps. One molecule of C_2H_5OH requires three molecules of O_2 for complete combustion. What mass of O_2 is required to react completely with 92.0 g of C_2H_5OH?

25. Bromine exists as two isotopes: Br-79 and Br-81. Calculate the relative abundance of each isotope. You will need to use information from the periodic table.

26. Examine the following double displacement reaction.

$$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

In this reaction, one formula unit of NaCl reacts with one formula unit of $AgNO_3$.
 (a) How many moles of NaCl react with one mole of $AgNO_3$?
 (b) What mass of $AgNO_3$ reacts with 29.2 g of NaCl?

27. The planet Zoltan is located in a solar system in the Andromeda galaxy. On Zoltan, the standard unit for the amount of substance is the wog and the standard unit for mass is the wibble. The Zoltanians, like us, chose carbon-12 to define their standard unit for the amount of substance. By definition, one wog of C-12 atoms contains 2.50×10^{21} atoms. It has a mass of exactly 12 wibbles.
 (a) What is the mass, in wibbles, of 1 wog of nitrogen atoms?
 (b) What is the mass, in wibbles, of 5.00×10^{-1} wogs of O_2?
 (c) What is the mass, in grams, of 1 wog of hydrogen atoms?

Communication

28. Use the definition of the Avogadro constant to explain why its value must be determined by experiment.

29. Why is carbon-12 the only isotope with an atomic mass that is a whole number?

30. Draw a concept map for the conversion of mass (g) of a sample to amount (mol) of a sample to number of molecules in a sample to number of atoms in a sample. Be sure to include proper units.

31. Explain why 1 mol of carbon dioxide contains 6.02×10^{23} molecules and not 6.02×10^{23} atoms.

Making Connections

32. The RNI (Recommended Nutrient Intake) of iron for women is listed as 14.8 mg per day. Ferrous gluconate, $Fe(C_6H_{11}O_7)_2$ is often used as an iron supplement for those who do not get enough iron in their diet because it is relatively easy for the body to absorb. Some iron-fortified breakfast cereals contain elemental iron metal as their source of iron.

(a) Calculate the number of moles of elemental iron, Fe, required by a woman, according to the RNI.

(b) What mass, in milligrams, of ferrous gluconate, would satisfy the RNI for iron?

(c) The term *bioavailability* refers to the extent that the body can absorb a certain vitamin or mineral supplement. There is evidence to suggest that the elemental iron in these iron-fortified cereals is absorbed only to a small extent. If this is the case, should cereal manufacturers be allowed to add elemental iron at all? How could cereal manufacturers assure that the consumer absorbs an appropriate amount of iron? Would adding more elemental iron be a good solution? List the pros and cons of adding more elemental iron, then propose an alternative solution.

33. Vitamin B_3, also known as niacin, helps maintain the normal function of the skin, nerves, and digestive system. The disease pellagra results from a severe niacin deficiency. People with pellagra experience mouth sores, skin irritation, and mental deterioration. Niacin has the following formula: $C_6H_5NO_2$. Often vitamin tablets contain vitamin B_3 in the form of niacinamide, $C_6H_6N_2O$, which is easier for the body to absorb.

(a) A vitamin supplement tablet contains 100 mg of niacinamide. What mass of niacin contains an equivalent number of moles as 100 mg of niacinamide?

(b) Do some research to find out how much niacin an average adult should ingest each day.

(c) Do some research to find out what kinds of food contain niacin.

(d) What are the consequences of ingesting too much niacin?

(e) Choose another vitamin to research. Find out its chemical formula, its associated recommended nutrient intake, and where it is found in our diet. Prepare a poster to communicate your findings.

Answers to Practice Problems and Short Answers to Section Review Questions

Practice Problems: **1.** 10.81 u **2.** 28.09 u **3.** 63.55 u **4.** 207.2 **5.** 99.8%, 0.2% **6.** very low, **7.** 72% **8.** 99.8%, 0.2% **9.** 3.37×10^{16} rows **10.** 2.80×10^{14} **11.** 1.44×10^{27} km **12.** 1.91×10^{16} a **13.** 1.07×10^{22} **14.** 2.83×10^{20} **15.** 1.3×10^{23} **16.** 3.35×10^{25} **17.(a)** 1.5×10^{24} **(b)** 2.1×10^{25} **(c)** 6.0×10^{24} **18.(a)** 4.99×10^{23} **(b)** 9.98×10^{23} **19.** 12.8 mol **20.** 1.33×10^{3} mol **21.** 1.84×10^{-2} mol **22.** 1.16×10^{-1} mol **23.(a)** 131.29 g/mol **(b)** 190.23 g/mol **(c)** 137.33 g/mol **(d)** 127.60 g/mol **24.(a)** 17.04 g/mol **(b)** 180.2 g/mol **(c)** 294.2 g/mol **(d)** 399.9 g/mol **25.** 183.68 g/mol **26.** 131.7 g/mol **27.(a)** 46.8 g **(b)** 1.20×10^{2} g **(c)** 1.05×10^{9} **(d)** 3.66×10^{-3} **28.(a)** 1.5 mol Cl_2 **(b)** 7.31 mol O_2 **29.** 1.00×10^{3} g **30.** 26.6 kg **31.(a)** 1.07 mol **(b)** 1.24×10^{2} mol **(c)** 14.2 mol **(d)** 7.75×10^{-4} mol **32.(a)** 0.652 mol **(b)** 0.156 mol **(c)** 1.76×10^{6} **(d)** 3.49×10^{-8} **33.** 1.7×10^{2} mol **34.** 1.75×10^{2} mol **35.(a)** 1.36×10^{3} g **(b)** 9.95 g **(c)** 7.32×10^{2} g **(d)** 2.15×10^{7} **36.** 2.93×10^{-21} g **37.** 7.95×10^{-23} g **38.** 6.25×10^{-22} g **39.(a)** 3.34×10^{23} molecules **(b)** 9.84×10^{23} molecules **(c)** 1.05×10^{23} molecules **(d)** 2.50×10^{20} formula units **40.** 4.2×10^{5} atoms **41.** 1.36×10^{20} **42.(a)** 2.14×10^{20} **(b)** 4.27×10^{20}

Section Review: **5.1:** **2.** 24.31 u **3.** No **5.2:** **2.** 1.00×10^{14} **3.** 1.9×10^{16} years **4.** 2.08×10^{24} atoms **5.(a)** 4.25 mol **(b)** 12.8 mol **6.** 3.0×10^{23} atoms **7.(a)** 3.43×10^{24} molecules **(b)** 2.06×10^{25} atoms **8.(a)** 7.04×10^{23} molecules **(b)** 3.52×10^{24} atoms **(c)** 2.11×10^{24} atoms **10.** Compound **5.3:** **2.(a)** 4.61 mol **(b)** 2.78×10^{24} **3.(a)** 4.67×10^{-23} g **(b)** 1.69×10^{25} u **4.(a)** 46.1 g **(b)** 4.75×10^{23} formula units **(c)** 9.50×10^{23} **5.** 5.01×10^{22} atoms **6.** 2.50×10^{-3} g