INVESTIGATIVE PHENOMENON

GO ONLINE to Engage with real-world phenomena by watching a video and to complete a CER interactive worksheet.

Why do we quantify matter in different ways?

Chemical Quantities

Avocados are typically sold by the count. Eggs and bread rolls are often packaged in groups of 12 we call a dozen. Some items, like apples and nuts, are sold by weight. Liquid items like milk are usually sold by volume. In chemistry, we quantify matter in the same way: by count, mass, and volume. Once you have viewed the Investigative Phenomenon video and used the claim-evidence-reasoning worksheet to draft an explanation, answer these reflection questions about quantifying matter.

1 **CCC Scale, Proportion, and Quantity** Estimate how long it would take you to count a dozen eggs. Now estimate how long it would take you to count the number of atoms in a single egg. If you knew how many atoms were in a single egg, how would you use that information to estimate the number of atoms in a dozen eggs?

2 **SEP Identify Unknowns** While eating from a can of mixed nuts, you wonder about the ratio of almonds to cashews. Describe how you would determine the percentage of almonds and the percentage of cashews in the can.

EXPERIENCE 1

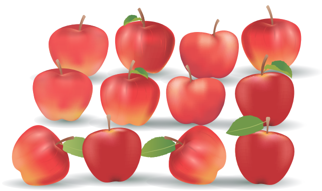
The Mole Concept

GO ONLINE to Explore and Explain moles and atomic mass.

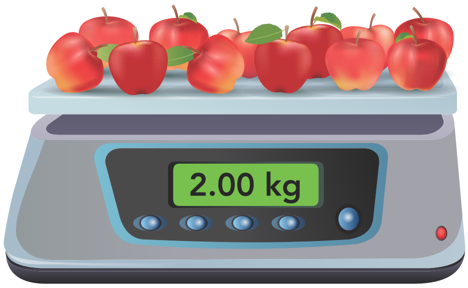
Measuring Matter

Chemistry is a quantitative science, which means we have to quantify how much matter we have. There are three basic ways to quantify matter: by count, by mass, and by volume. Knowing how the count, mass, and volume of an item relate to a common unit allows you to convert among these units.

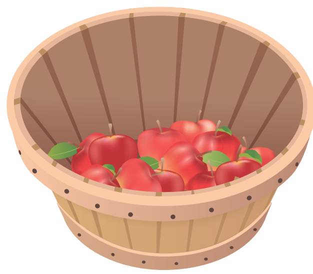
Converting Among Units Apples can be quantified by their count, volume, or mass. You can use conversion factors to convert between these quantities of apples.



Count
1 dozen apples = 12 apples



Mass
1 dozen apples = 2.0 kg



Volume
1 dozen apples = 0.20 bushels

$$\frac{1 \text{ dozen apples}}{12 \text{ apples}} \qquad \frac{1 \text{ dozen apples}}{2.0 \text{ kg apples}} \qquad \frac{1 \text{ dozen apples}}{0.20 \text{ bushel apples}}$$

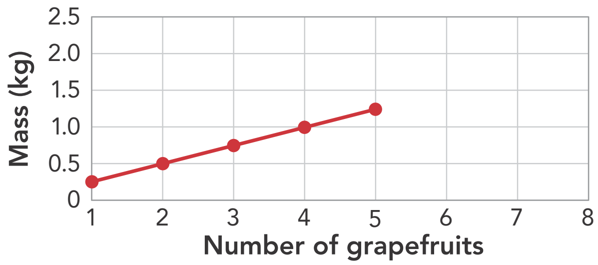
These conversion factors relate the count, mass, and volume of apples. The conversion factors for count and mass can be used to calculate the mass of one apple.

Mass of One Apple

$$1 \cancel{\text{apple}} \times \frac{1 \cancel{\text{dozen apples}}}{12 \cancel{\text{apples}}} \times \frac{2.0 \text{ kg apples}}{1 \cancel{\text{dozen apples}}} = 0.17 \text{ kg}$$

In order for the units to cancel, dozen apples must be in the denominator.

(3) **SEP Interpret Data** The graph shows the total mass as a function of the number of grapefruits. What is the slope of the best-fit line? What does this value mean? Predict the mass of 8 grapefruits.



SAMPLE PROBLEM

Finding Mass From a Count

What is the mass of 90 apples if 1 dozen of the apples has a mass of 2.0 kg?

ANALYZE List the knowns and the unknown.

Knowns	Unknown
number of apples = 90	mass of 90 apples = ? kg
12 apples = 1 dozen apples	
1 dozen apples = 2.0 kg apples	

CALCULATE Solve for the unknown.

Identify the steps to convert from number, or count, to mass.

number of apples → dozens of apples → mass of apples

Multiply the number of apples by the two conversion factors needed to convert from number of apples to mass of apples.

$$90 \cancel{\text{apples}} \times \frac{1 \cancel{\text{dozen apples}}}{12 \cancel{\text{apples}}} \times \frac{2.0 \text{ kg apples}}{1 \cancel{\text{dozen apples}}} = \boxed{15 \text{ kg apples}}$$

EVALUATE Does the result make sense?

A dozen apples has a mass of 2.0 kg, and 90 apples is less than 10 dozen apples, so the mass should be less than 20 kg of apples (10 dozen × 2.0 kg/dozen).

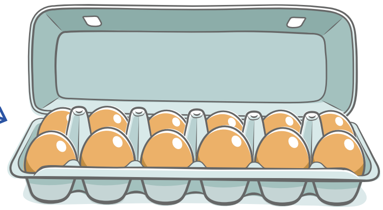
(4) **SEP Use Mathematics** Assume 2.0 kg of apples is 1 dozen and that each apple has 8 seeds. How many seeds are in 14 kg of apples?

GO ONLINE for more practice problems.

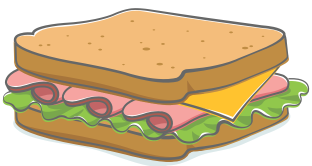
Counting With Moles

Counting the number of atoms in a gram of substance would take a very long time. To simplify the counting of atoms, chemists group them into a specified number called a mole. A **mole** (mol) of a substance is 6.02×10^{23} representative particles of that substance. The term **representative particle** refers to the species present in a substance, usually atoms, molecules, or formula units. In compounds, each molecule or formula unit is a representative particle, but they are made of multiple elements. For example, in every water molecule there is one oxygen atom and two hydrogen atoms. The number of representative particles in a mole, 6.02×10^{23}, is called **Avogadro's number** after the Italian scientist Amedeo Avogadro.

Representative Particles Representative particles are usually atoms, molecules, or formula units. Other items, such as eggs or sandwiches, can also be representative particles.



Grouping eggs in a dozen is similar to grouping atoms where each atom is a representative particle. One egg is the representative particle in a dozen eggs.



This entire sandwich is similar to a representative particle of a compound because it is made of certain amounts of each ingredient. One sandwich is a representative particle made of 2 slices of bread, 3 pieces of ham, 1 slice of cheese, and 1 leaf of lettuce. To make a dozen identical sandwiches, you need 24 slices of bread. To make a mole of identical sandwiches, you need 1.20×10^{24} slices of bread.

5 **SEP Develop Models** **The table shows examples of representative particles for several types of substances. Develop a mental model for the concept of representative particles and use your model to complete the table.**

Representative Particles and Moles

Substance	Representative Particle	Chemical Formula	Representative Particles in 1.00 mol
Copper	Atom	Cu	6.02×10^{23}
Atomic nitrogen			
Nitrogen gas	Molecule	N_2	6.02×10^{23}
Water			
Calcium fluoride	Formula unit	CaF_2	6.02×10^{23}
Sodium chloride			

SAMPLE PROBLEM

Converting Number of Atoms to Moles

Magnesium is a light metal used in the manufacture of aircraft and tools. How many moles of magnesium is 1.25×10^{23} atoms of magnesium?

ANALYZE List the knowns and the unknown.

Knowns	Unknown
number of atoms = 1.25×10^{23} atoms Mg	moles = ? mol Mg

CALCULATE Solve for the unknown.

Identify the relationship between moles and number of representative particles.

$$1 \text{ mol Mg} = 6.02 \times 10^{23} \text{ atoms Mg}$$

Write the conversion factors based on this relationship.

$$\frac{1 \text{ mol Mg}}{6.02 \times 10^{23} \text{ atoms Mg}} \text{ and } \frac{6.02 \times 10^{23} \text{ atoms Mg}}{1 \text{ mol Mg}}$$

Identify the conversion factor needed to convert from atoms to moles.

$$\frac{1 \text{ mol Mg}}{6.02 \times 10^{23} \text{ atoms Mg}}$$

The units in the known should be in the denominator so that the units will cancel.

Multiply the number of atoms of Mg by the conversion factor.

$$1.25 \times 10^{23} \cancel{\text{atoms Mg}} \times \frac{1 \text{ mol Mg}}{6.02 \times 10^{23} \cancel{\text{atoms Mg}}}$$

$$= \boxed{0.208 \text{ mol Mg}}$$

EVALUATE Does the result make sense?

The given number of atoms is less than one fourth of Avogadro's number, so the answer should be less than one fourth (0.25) mol of atoms. The answer should have three significant figures.

6 **SEP Use Math** How many moles is 2.17×10^{23} representative particles of bromine gas?

7 **SEP Use Math** How many moles is 2.80×10^{24} atoms of silicon?

GO ONLINE for more practice problems.

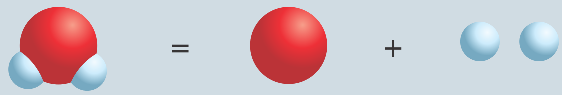
Volume and Mass of a Mole

A dozen roses, a dozen bagels, and a dozen eggs all have 12 representative particles. Because roses, bagels, and eggs are different representative particles, their respective masses and volumes are different. Similar to the number of representative particles in a dozen, a mole is 6.02×10^{23} representative particles of a substance, no matter what the substance.

Representative particles can come in different masses and sizes. Therefore, a mole of a substance will have a mass and volume dependent on the properties of the representative particle.

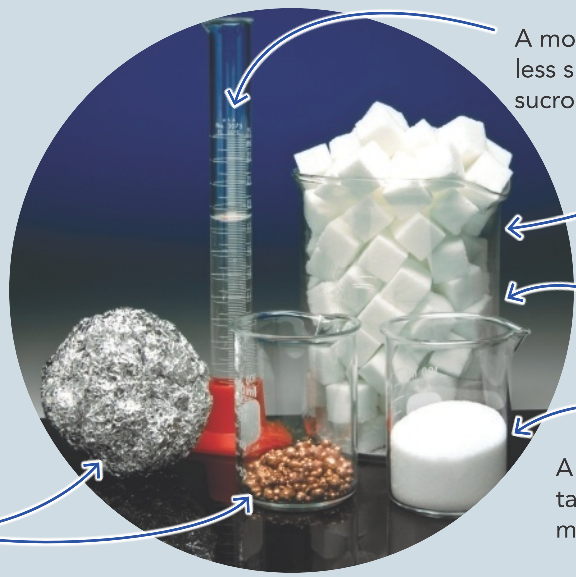
A Mole of Common Substances
Common substances may have atoms, molecules, or formula units as their representative particles. A water molecule is the representative unit of water. Aluminum and copper both have atoms as their representative particles. The volume of these substances depends on the properties of these particles.

A water molecule is composed of one oxygen atom and two hydrogen atoms.



Therefore, 1 mole of water has 3 moles of atoms.

1 mole of H_2O = 1 mole of O + 2 moles of H



A mole of water takes up less space than a mole of sucrose (sugar, $C_{12}H_{22}O_{11}$).

A mole of copper takes up less space than a mole of aluminum because copper is denser than aluminum.

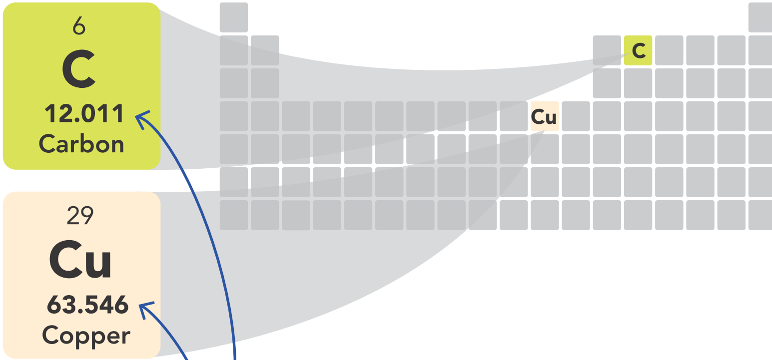
A mole of table salt (NaCl) takes up less space than a mole of sugar ($C_{12}H_{22}O_{11}$).

A mole of water has less mass than a mole of sugar or a mole of copper. The mass of 1 mole of a substance is called the **molar mass** and is expressed in units of g/mol. Similarly, the atomic mass is the mass of a mole of the element expressed in grams. Therefore, the molar mass of elements can be read directly from the periodic table.

Molar Mass

How do you obtain the **molar mass from the periodic table?**

Atomic Weight The periodic table gives the atomic weight of each element. Atomic weight is a dimensionless quantity that represents the **average mass of atoms of an element divided by 1/12 the atomic mass of a carbon-12 atom.**

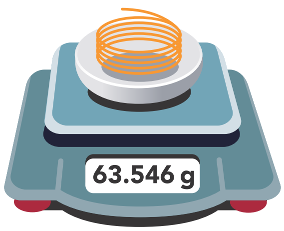


Molar Mass The **molar mass is the mass of 1 mole of an element.** It is the atomic weight multiplied by 1 g/mol. To obtain the molar mass, read the atomic weight from the periodic table and add the unit g/mol.

1 mole of carbon

Molar mass of carbon
= 12.011 g/mol

1 mole of copper



Molar mass of copper
= 63.546 g/mol

Molar Mass as Conversion Factor
The **molar mass can be represented as a fraction and used as a conversion factor** to convert a known mass into moles or known moles into mass.
The molar mass is often rounded to 1 decimal place to simplify calculations.

$$\frac{12.0 \text{ g C}}{1 \text{ mol C}} \text{ or } \frac{1 \text{ mol C}}{12.0 \text{ g C}}$$

$$\frac{63.5 \text{ g Cu}}{1 \text{ mol Cu}} \text{ or } \frac{1 \text{ mol Cu}}{63.5 \text{ g Cu}}$$

8 **CCC Scale, Proportion, and Quantity** The table shows how many moles are in 6 grams of four elements. The equation shows how to use carbon's molar mass to find the moles of carbon. Complete the table.

$$6.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.0 \text{ g C}} = 0.50 \text{ mol C}$$

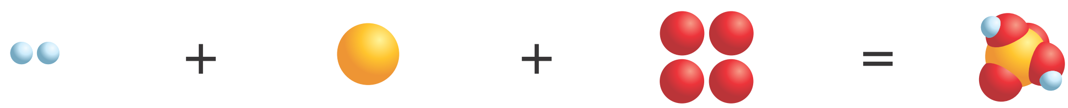
Converting Mass to Moles

Substance	Mass (g)	Molar Mass (g/mol)	Moles (mol)
Carbon	6.0	12.0	0.50
Sulfur	6.0		
Aluminum	6.0		
Boron	6.0		

Molar Mass of Compounds

The molar mass of a compound depends on the number of moles and the molar mass of each element in the compound's representative particle. The number of atoms of each element in the chemical formula is also the number of moles of that element in one mole of the compound. Therefore, the molar masses of the elements in the compound are used to determine the molar mass of the compound. For example, you can use the molar masses of sulfur, oxygen, and hydrogen to determine the molar mass of sulfuric acid (H_2SO_4).

Molar Mass of Sulfuric Acid Adding the masses of 2 moles of hydrogen, 1 mole of sulfur, and 4 moles of oxygen gives the molar mass of H_2SO_4.



2 H atoms + 1 S atom + 4 O atoms = 1 H_2SO_4 molecule

The moles of H, S, and O in 1 mol H_2SO_4 are the same as the number of atoms of each element in one molecule of H_2SO_4.

2 mol H + 1 mol S + 4 mol O = 1 mol H_2SO_4

1	1.008	H	Hydrogen
16	32.06	S	Sulfur
8	15.999	O	Oxygen

Round the values in the periodic table to one decimal place to simplify the molar mass calculation.

Multiply the moles of each element by its molar mass.

$$\left(2\ \cancel{\text{mol H}} \times \frac{1.0\text{ g H}}{1\ \cancel{\text{mol H}}}\right) + \left(1\ \cancel{\text{mol S}} \times \frac{32.1\text{ g S}}{1\ \cancel{\text{mol S}}}\right) + \left(4\ \cancel{\text{mol O}} \times \frac{16.0\text{ g O}}{1\ \cancel{\text{mol O}}}\right) = 98.1\text{ g/mol}$$

9 **CCC Scale, Proportion, and Quantity** The table shows the total mass of carbon dioxide (CO_2) as a function of the number of moles of CO_2. Complete the table.

Mass of CO_2			
CO_2 (mol)	C (mol)	O (mol)	Total mass CO_2 (g)
1	1	2	44.0
2			
3			
4			

SAMPLE PROBLEM

Find the Molar Mass of a Compound

The decomposition reaction of hydrogen peroxide (H_2O_2) provides sufficient energy to launch a rocket. What is the molar mass of hydrogen peroxide?

ANALYZE List the knowns and the unknown.

Knowns	Unknown
molecular formula = H_2O_2	molar mass = ? g/mol
mass of 1 mol H = 1.0 g H	
mass of 1 mol O = 16.0 g O	

CALCULATE Solve for the unknown.

Convert moles of atoms to grams using conversion factors (g/mol) based on the molar mass of each element.

Multiply the molar mass of each element by its subscript in the chemical formula to find the total mass of the element.

$$2\ \cancel{\text{mol H}} \times \frac{1.0\ \text{g H}}{1\ \cancel{\text{mol H}}} = 2.0\ \text{g H}$$

$$2\ \cancel{\text{mol O}} \times \frac{16.0\ \text{g O}}{1\ \cancel{\text{mol O}}} = 32.0\ \text{g O}$$

Add the masses of the elements to determine the molar mass of the compound.

mass of 1 mol H_2O_2 = 2.0 g H + 32.0 g O = 34.0 g

molar mass of H_2O_2 = 34.0 g/mol

EVALUATE Does the result make sense?

There are two hydrogen atoms and two oxygen atoms in H_2O_2, so the answer should be two times the molar mass of hydrogen and oxygen. The answer is expressed to the tenths place because the numbers being added are expressed to the tenths place.

10 **SEP Use Mathematics** Phosphorus trichloride (PCl_3) is used as a reactant in the production of many other chemicals that contain phosphorus, including pesticides and herbicides. Find the molar mass of PCl_3.

11 **SEP Use Mathematics** Baking soda is also known as sodium hydrogen carbonate. What is the mass of 1.00 mol of sodium hydrogen carbonate ($NaHCO_3$)?

GO ONLINE for more practice problems.

Revisit

INVESTIGATIVE PHENOMENON

GO ONLINE to Elaborate on and Evaluate your knowledge of moles as a counting unit by completing the discussion and writing activities.

In the CER worksheet you completed at the beginning of the investigation, you drafted an explanation for the ways in which matter is quantified. With a partner, reevaluate your arguments.

12 **SEP Construct an Explanation** A mole of table sugar (sucrose, $C_{12}H_{22}O_{11}$) is 342.3 g. A mole of table salt (sodium chloride, NaCl) is 58.45 g. What does the difference in macroscopic size of a mole of sugar and a mole of salt suggest about their microscopic representative particles? Considering the size of a mole of these materials, why do you think chemists use the concept of moles when dealing with macroscopic quantities?

Molar Relationships

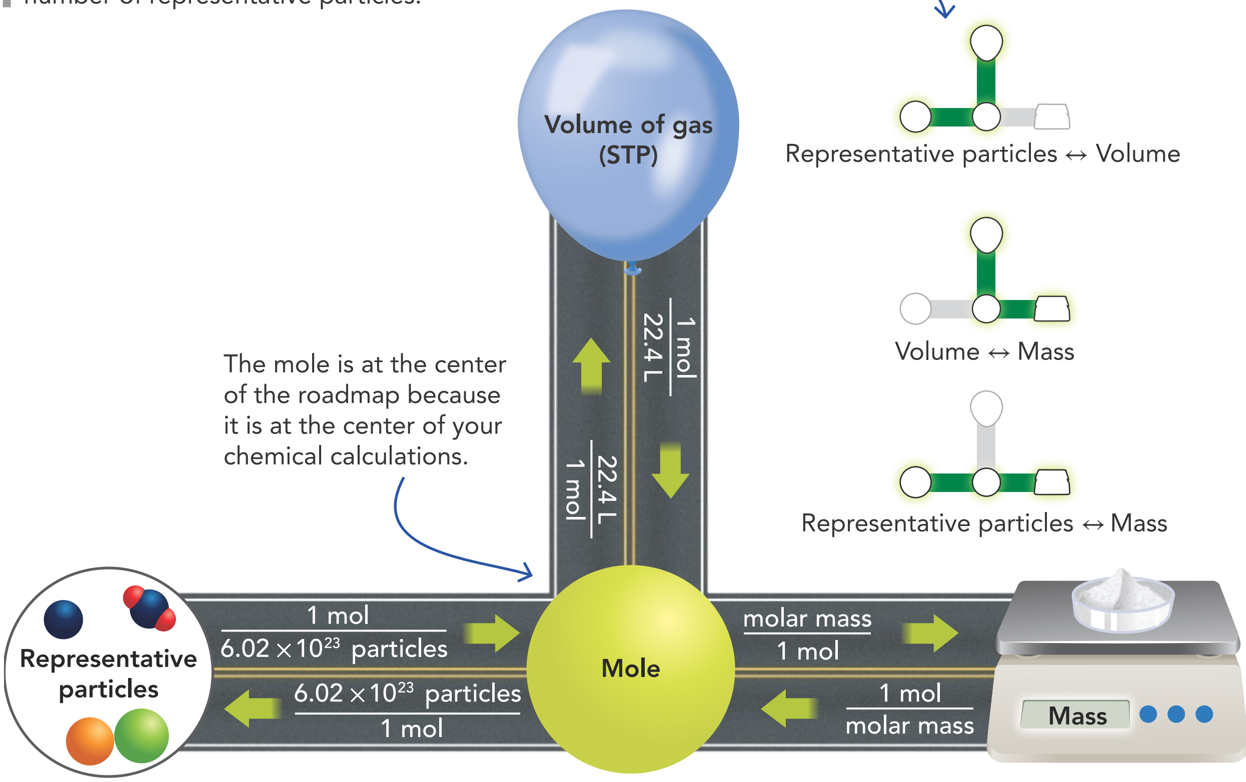
GO ONLINE to Explore and Explain how the mole relates to mass, volume, and the number of particles.

The Mole Roadmap

When you convert between mass, volume, and representative particles, you must use the mole as an intermediate step. The mole roadmap provides the path to convert one quantity to another, just as a real roadmap shows you how to get from Point A to Point B. It shows how the conversion factor you need to use depends on what you know and what you want to calculate. Including the appropriate units in the conversion factors will help keep you on the right path in your calculations. You know you have chosen the correct conversion factor(s) if the units cancel to give the desired unit for the answer.

The Mole Roadmap Follow the arrows on the roadmap to determine which conversion factors you need to use to convert between mass, volume, and number of representative particles.

You can choose different routes on the roadmap to get from one chemical quantity to another.



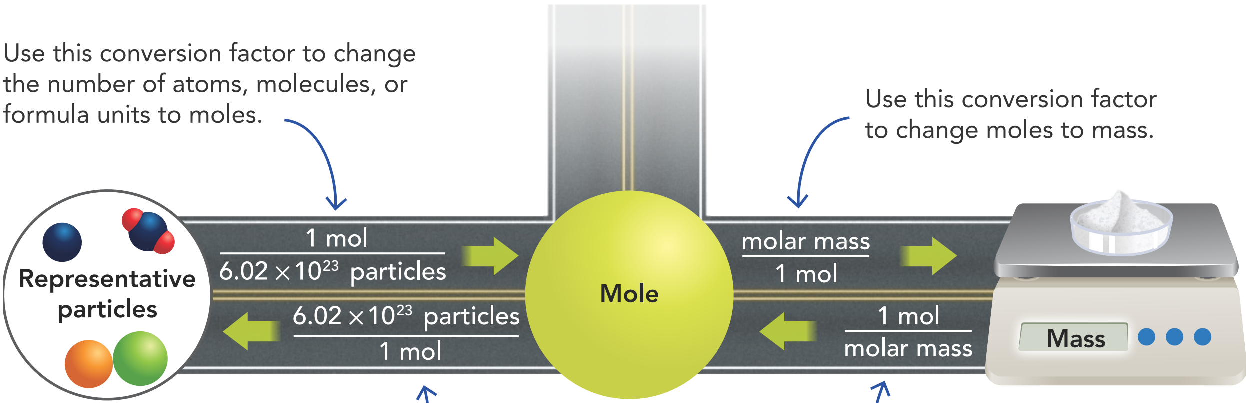
The Mole-Mass Relationship

The molar mass of a substance can be used to convert between the number of moles of a substance and its mass. The conversion factors for these calculations are based on this relationship: molar mass = 1 mol. The molar mass can be used as a conversion factor for individual elements or compounds.

Converting Between Mass, Moles, and Particles The mole-mass relationship can be used to convert from mass to moles or from moles to mass. The relationship between the mole and Avogadro's number can also be used to convert between representative particles and mass.

Use this conversion factor to change the number of atoms, molecules, or formula units to moles.

Use this conversion factor to change moles to mass.



Use this conversion factor to change moles to the number of atoms, molecules, or formula units.

Use this conversion factor to change mass to moles.

Applying the Mole-to-Mass Roadmap What is the mass, in grams, of 3.2 mol nitrogen dioxide (NO_2)?

Use the mole roadmap to determine how to convert moles to mass. → moles of $NO_2 \times \frac{\text{molar mass } NO_2}{1 \text{ mol } NO_2}$ ⇒ mass of NO_2

Use the subscripts from the chemical formula to find the mass of each element in the compound.

$$1 \cancel{\text{mol N}} \times \frac{14.0 \text{ g N}}{1 \cancel{\text{mol N}}} \Rightarrow 14.0 \text{ g N}$$

$$2 \cancel{\text{mol O}} \times \frac{16.0 \text{ g O}}{1 \cancel{\text{mol O}}} \Rightarrow 32.0 \text{ g O}$$

Add the mass of the elements to find the mass of 1 mole of the compound.

$$1 \text{ mol } NO_2 = 14.0 \text{ g N} + 32.0 \text{ g O}$$
$$= 46.0 \text{ g } NO_2$$

Multiply the given number of moles by the conversion factor. →

$$3.2 \cancel{\text{mol } NO_2} \times \frac{46.0 \text{ g } NO_2}{1 \cancel{\text{mol } NO_2}} \Rightarrow 147.2 \text{ g } NO_2$$

SAMPLE PROBLEM

Converting Moles to Mass

Items made of aluminum, such as aircraft parts and cookware, are resistant to corrosion because the aluminum reacts with oxygen in the air to form a coating of aluminum oxide (Al_2O_3). This tough, resistant coating prevents any further corrosion. What is the mass, in grams, of 9.45 mol aluminum oxide?

ANALYZE List the knowns and the unknown.

Knowns	Unknown
number of moles = 9.45 mol Al_2O_3	mass = ? g Al_2O_3

CALCULATE Solve for the unknown.

Identify the necessary conversion.

moles → mass

Multiply the moles by the molar masses to determine the mass of each element in the compound.

$$2\ \cancel{\text{mol Al}} \times \frac{27.0\ \text{g Al}}{1\ \cancel{\text{mol Al}}} = 54.0\ \text{g Al}$$

$$3\ \cancel{\text{mol O}} \times \frac{16.0\ \text{g O}}{1\ \cancel{\text{mol O}}} = 48.0\ \text{g O}$$

Use the subscripts to determine the moles of each element in the compound.

Add the mass of the elements to find the mass of 1 mole of the compound.

$$1\ \text{mol } Al_2O_3 = 54.0\ \text{g Al} + 48.0\ \text{g O} = 102.0\ \text{g } Al_2O_3$$

Identify the conversion factor for converting from moles to mass.

$$\frac{102.0\ \text{g } Al_2O_3}{1\ \text{mol } Al_2O_3}$$

The known unit (mol) is in the denominator and the unknown unit (g) is in the numerator.

Multiply the given number of moles by the conversion factor.

$$9.45\ \cancel{\text{mol } Al_2O_3} \times \frac{102.0\ \text{g } Al_2O_3}{1\ \cancel{\text{mol } Al_2O_3}} = \boxed{964\ \text{g } Al_2O_3}$$

EVALUATE Does the result make sense?

The number of moles of Al_2O_3 is approximately 10, and each mole has a mass of approximately 100 g. The answer should be close to 1000 g. The answer has been rounded to the correct number of significant figures.

13 **SEP Use Mathematics** Calculate the mass, in grams, of 2.50 mol of iron(II) hydroxide.

GO ONLINE for more practice problems.

SAMPLE PROBLEM

Converting Mass to Moles

When iron is exposed to air, it undergoes a corrosion reaction to form a red-brown rust. Rust is iron(III) oxide (Fe_2O_3). How many moles of iron(III) oxide are contained in 92.2 g of pure Fe_2O_3?

ANALYZE **List the knowns and the unknown.**

Knowns	Unknown
mass = 92.2 g Fe_2O_3	number of moles = ? mol Fe_2O_3

CALCULATE **Solve for the unknown.**

Identify the necessary conversion.

mass → moles

Use the subscripts to determine the moles of each element in the compound. Multiply the moles by the molar masses to determine the mass of each element in the compound.

$$2\ \cancel{\text{mol Fe}} \times \frac{55.8\ \text{g Fe}}{1\ \cancel{\text{mol Fe}}} = 111.6\ \text{g Fe}$$

$$3\ \cancel{\text{mol O}} \times \frac{16.0\ \text{g O}}{1\ \cancel{\text{mol O}}} = 48.0\ \text{g O}$$

Add the mass of the elements to find the mass of 1 mole of the compound.

$$1\ \text{mol Fe}_2\text{O}_3 = 111.6\ \text{g Al} + 48.0\ \text{g O}$$
$$= 159.6\ \text{g Fe}_2\text{O}_3$$

Identify the conversion factor for converting mass to moles.

$$\frac{1\ \text{mol Fe}_2\text{O}_3}{159.6\ \text{g Fe}_2\text{O}_3}$$

Note that the known unit (g) is in the denominator and the unknown unit (mol) is in the numerator.

Multiply the given mass by the conversion factor.

$$92.2\ \cancel{\text{g Fe}_2\text{O}_3} \times \frac{1\ \text{mol Fe}_2\text{O}_3}{159.6\ \cancel{\text{g Fe}_2\text{O}_3}} = \boxed{0.578\ \text{mol Fe}_2\text{O}_3}$$

EVALUATE **Does the result make sense?**

The given mass is slightly larger than the mass of one-half mole of Fe_2O_3, so the answer should be slightly larger than one-half mol.

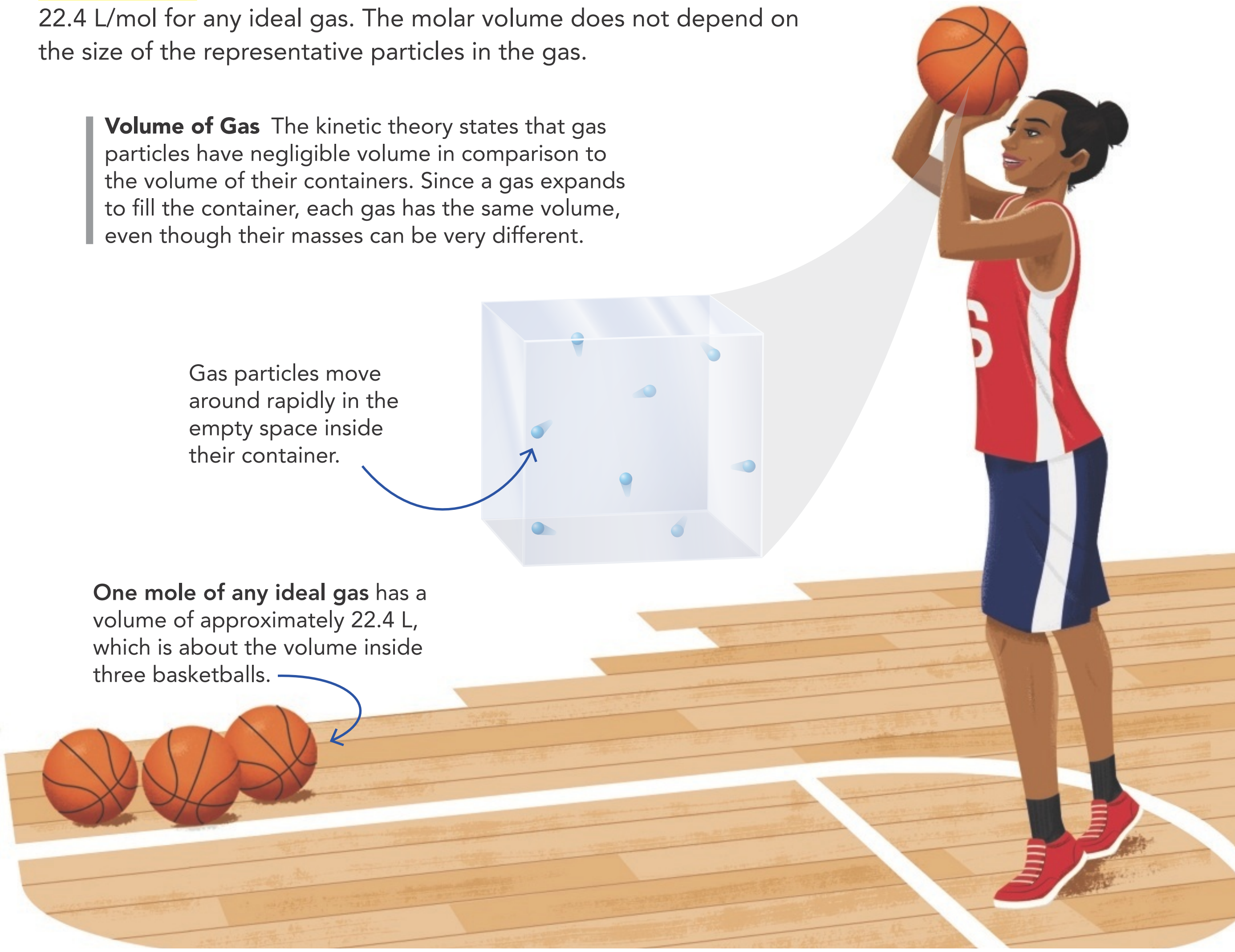
14 **SEP Use Mathematics** **Calculate the number of moles in 75.0 g of dinitrogen trioxide.**

GO ONLINE for more practice problems.

Avogadro's Hypothesis

Avogadro's hypothesis states that any two samples of gas with an equal number of particles will have the same volume when they are held at the same pressure and temperature. It was proposed in 1811 by Amedeo Avogadro. The volume of a gas is usually measured at standard temperature and pressure (STP)—0°C and 101.3 kPa, or 1 atm. The **molar volume** is the volume of 1 mole of a gas at STP and is 22.4 L/mol for any ideal gas. The molar volume does not depend on the size of the representative particles in the gas.

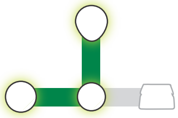
Volume of Gas The kinetic theory states that gas particles have negligible volume in comparison to the volume of their containers. Since a gas expands to fill the container, each gas has the same volume, even though their masses can be very different.



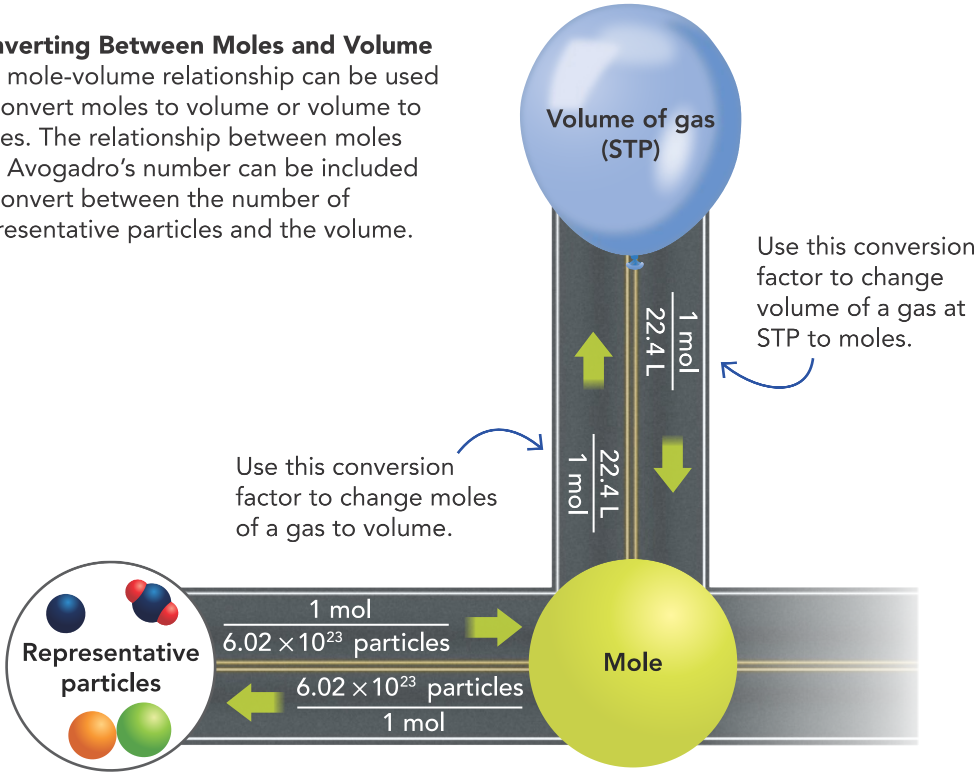
15 **SEP Argue From Evidence** Imagine a basketball filled with liquid water. Would the basketball still hold the same number of molecules? Can you use 22.4 L/mol as the molar volume for liquid water?

The Mole-Volume Relationship

The molar volume of a gas is used to convert between the number of moles of a gas and the volume of a gas. The conversion factors are based on this relationship: molar volume = 1 mol. The molar volume can be used as a conversion factor for individual gaseous elements or compounds.



Converting Between Moles and Volume The mole-volume relationship can be used to convert moles to volume or volume to moles. The relationship between moles and Avogadro's number can be included to convert between the number of representative particles and the volume.



16 **SEP Analyze Data** The table shows experimental data for the volume as a function of the number of moles for nitrous oxide (N_2O) gas, a sedative known as laughing gas. The slope of the best-fit line for these data can be used to convert volume to moles. Graph the volume as a function of moles, draw a best-fit line on the graph, and determine its slope.

Volume and Moles of N_2O

Moles (mol)	Volume (L)
0.15	3.4
0.25	5.6
0.35	7.8
0.45	10.1
0.55	12.3

SAMPLE PROBLEM

Calculating Gas Quantities at STP

Sulfur dioxide (SO_2) is a gas produced by burning coal. It is an air pollutant and one of the causes of acid rain. Determine the volume, in liters, of 0.60 mol SO_2 gas at STP.

ANALYZE List the knowns and the unknown.

Knowns	Unknown
number of moles = 0.60 mol SO_2	volume = ? L SO_2
1 mol SO_2 = 22.4 L SO_2 at STP	

CALCULATE Solve for the unknown.

Identify the conversion factor relating moles of SO_2 gas to volume of SO_2 gas at STP.

$$\frac{22.4\ \text{L}\ SO_2}{1\ \text{mol}\ SO_2}$$

Multiply the given number of moles by the conversion factor.

$$0.60\ \cancel{\text{mol}\ SO_2} \times \frac{22.4\ \text{L}\ SO_2}{1\ \cancel{\text{mol}\ SO_2}} = \boxed{13\ \text{L}\ SO_2}$$

EVALUATE Does the result make sense?

One mole of any gas at STP has a volume of 22.4 L, so 0.60 mol should have a volume slightly larger than one-half of this value. The answer should have two significant figures.

17 **SEP Use Mathematics** At STP, how many moles are in the given volumes of the following gases?

a. 67.2 L SO_2

b. 0.880 L He

c. 1.00×10^3 L C_2H_6

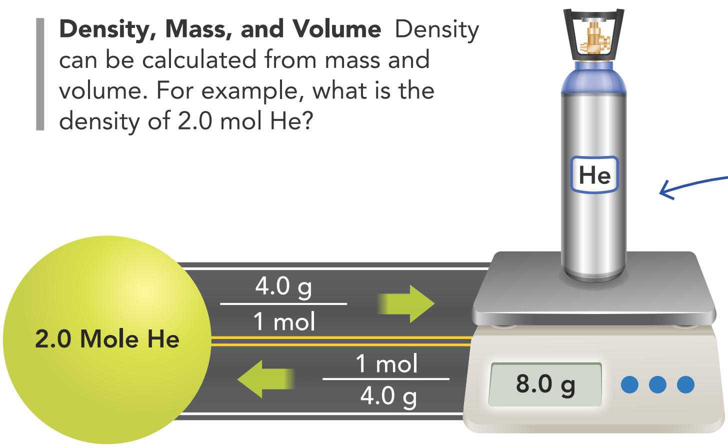
GO ONLINE for more practice problems.

Mass and Density

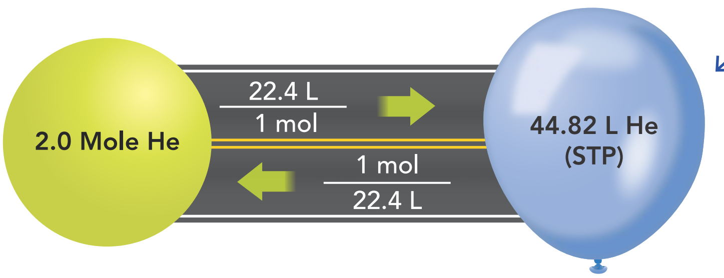
A gas-filled balloon will either sink or float in the air depending on whether the density of the gas inside the balloon is greater or less than the density of the surrounding air. **Density** is the ratio of the mass of a substance to its volume. Although the molar volumes of gases are approximately the same (22.4 L/mol) at STP, different gases will have different densities. The density of a gas at STP can be determined using one of the following expressions:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad \text{density} = \frac{\text{molar mass}}{\text{molar volume}}$$

Density, Mass, and Volume Density can be calculated from mass and volume. For example, what is the density of 2.0 mol He?



The **mole-mass relationship** can be used to determine the mass of 2.0 mol of He gas.



The **mole-volume relationship** can be used to determine the volume of 2.0 mol He gas.

$$\text{density He} = \frac{\text{mass}}{\text{volume}} = \frac{8.0 \text{ g He}}{44.82 \text{ L He}} = 0.18 \text{ g/L He}$$

Helium has a density of 0.18 g/L at STP, which is less than the density of air (1.23 g/L). Therefore, a helium-filled balloon floats.

$$\text{density He} = \frac{\text{molar mass He}}{\text{molar volume He}} = \frac{\dfrac{4.0 \text{ g He}}{\cancel{1 \text{ mol He}}}}{\dfrac{22.4 \text{ L He}}{\cancel{1 \text{ mol He}}}} = 0.18 \text{ g/L He}$$

Since density is a ratio, you can also calculate it directly from the molar mass and molar volume. The moles cancel!

18 **SEP Interpret Data** The table shows several gases, their molar masses, and their densities at STP. Complete the table by determining whether or not a balloon filled with the gas will float in air. The density of air is 1.23 g/L.

Density of Gases

Gas	Molar Mass (g/mol)	Density (g/L)	Does It Float?
He	4.0	0.18	Yes
CH_4	16.0	0.71	
CO_2	44.0	1.96	
C_6H_6	78.1		
NH_3	17.0		

Revisit INVESTIGATIVE PHENOMENON

GO ONLINE to Elaborate on and Evaluate your knowledge of mole-mass and mole-volume relationships by completing the peer review and data analysis activities.

In the CER worksheet you completed at the beginning of the investigation, you drafted an explanation for the ways in which matter is quantified. With a partner, reevaluate your arguments.

19 **SEP Design a Solution** A simple bread recipe calls for 400 g of flour, 7 g of salt (NaCl), 1 g of yeast, and 0.3 L of water (H_2O). You have exactly 0.6 mol of salt. If you want to use all of the salt, how many loaves of bread could you make? How much of each of the other ingredients would you need?

Percent Composition and Empirical Formulas

GO ONLINE to Explore and Explain percent composition and molecular formulas.

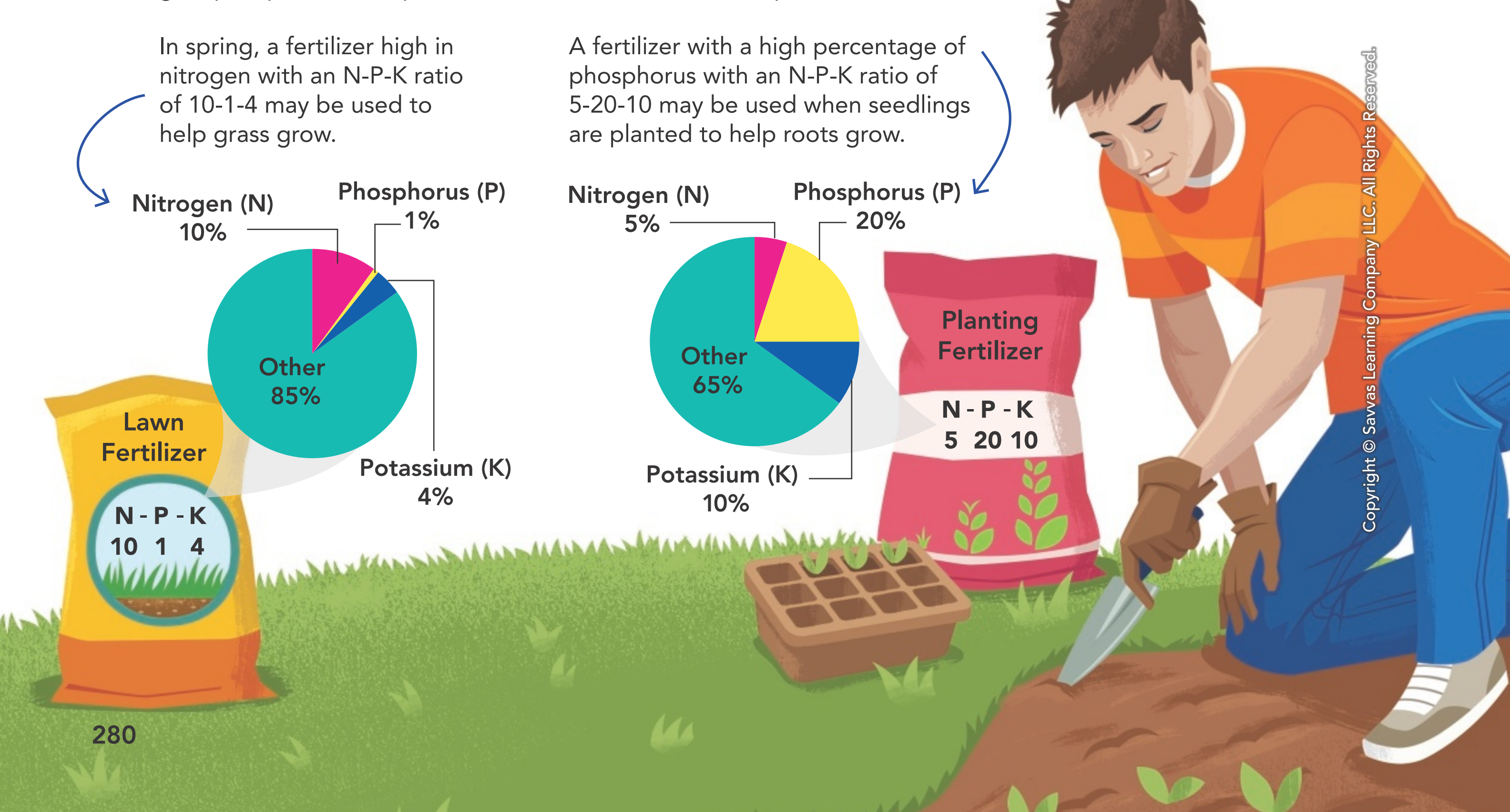
Percent Composition of a Compound

In a given compound, the ratio of the atoms of each element is fixed. Therefore, the ratio of the masses of the elements is also fixed. The **law of definite proportions** states that a compound contains its component elements in a fixed ratio by mass. For example, in carbon dioxide (CO_2), the ratio of C:O is 1 mol:2 mol. The ratio of the mass is 12.0 g C:32.0 g O, or 0.38 g C:1.0 g O. This ratio is the same for any amount of CO_2.

The relative amounts of the elements in the compound are expressed as the **percent composition,** or the percent by mass of each element in the compound. The mathematical expression for the percent by mass of an element in a compound is as follows:

$$\text{\% by mass of element} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\%$$

Fertilizer The relative amounts, or the percent composition, of nitrogen, phosphorus, and potassium (N-P-K) in fertilizer is important.



20 **SEP Plan an Investigation** Plan an investigation to determine the percentage by mass of cashews and almonds in a jar of mixed nuts.

SAMPLE PROBLEM

Percent Composition From Mass Data

When a 13.60-g sample of a compound containing only magnesium and oxygen is decomposed, 5.40 g of oxygen is obtained. What is the percent composition of this compound?

ANALYZE **List the knowns and the unknowns.**

Knowns	Unknowns
mass of compound = 13.60 g	percent by mass of Mg = ?% Mg
mass of oxygen = 5.40 g O	percent by mass of O = ?% O
mass of magnesium = 13.6 g − 5.40 g O = 8.20 g Mg	

CALCULATE **Solve for the unknowns.**

Determine the percent by mass of Mg in the compound.

$$\%\text{Mg} = \frac{\text{mass of Mg}}{\text{mass of compound}} \times 100\%$$

$$= \frac{8.20\ \cancel{g}}{13.60\ \cancel{g}} \times 100\% = 60.3\%\ \text{Mg}$$

Determine the percent by mass of O in the compound.

$$\%\text{O} = \frac{\text{mass of O}}{\text{mass of compound}} \times 100\%$$

$$= \frac{5.40\ \cancel{g}}{13.60\ \cancel{g}} \times 100\% = 39.7\%\ \text{O}$$

EVALUATE **Does the result make sense?**

The percents of the elements add up to 100%.

21 **SEP Use Mathematics** When a 14.2-g sample of mercury(II) oxide is decomposed into its elements by heating, 13.2 g Hg is obtained. What is the percent composition of this compound?

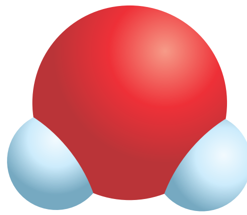
GO ONLINE for more practice problems.

Percent Composition From Chemical Formulas

The **law of constant composition** states that any sample of a compound will be made up of the same elements in the same ratio. Because of this, you can calculate the percent composition of a compound using its chemical formula. The mathematical expression for the percent by mass of an element in a compound is as follows:

$$\% \text{ by mass of element} = \frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100\%$$

Percent Composition of Water Any sample of pure water will have the same proportions of hydrogen and oxygen. To determine the percent composition of each element in water, calculate its proportion by mass.



Step 1 Use the subscripts to determine how many moles of each element are in the compound.

$$1 \text{ mol } H_2O = 2 \text{ mol H} + 1 \text{ mol O}$$

Step 2 Convert moles to mass for each element in the compound.

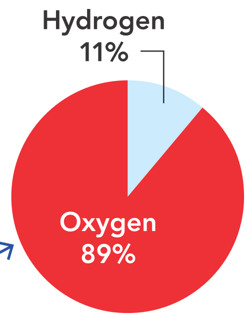
$$2\ \cancel{\text{mol H}} \times \boxed{\frac{1.0 \text{ g H}}{1\ \cancel{\text{mol H}}}} \rightarrow 2.0 \text{ g H} \qquad 1\ \cancel{\text{mol O}} \times \boxed{\frac{16.0 \text{ g O}}{1\ \cancel{\text{mol O}}}} \rightarrow 16.0 \text{ g O}$$

Step 3 Calculate the percent composition of each element.

$$\% \text{ by mass of element} \rightarrow \boxed{\frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}}} \times 100\%$$

$$\% \text{ by mass of H} \rightarrow \boxed{\frac{2.0\ \cancel{g}}{18.0\ \cancel{g}}} \times 100\% \rightarrow 11\% \text{ H}$$

$$\% \text{ by mass of O} \rightarrow \boxed{\frac{16.0\ \cancel{g}}{18.0\ \cancel{g}}} \times 100\% \rightarrow 89\% \text{ O}$$



Any sample of pure water is 11% hydrogen to 89% oxygen by mass.

22 **SEP Use Mathematics** In the "heavy water" that is used in nuclear reactors, the hydrogen in water is replaced with deuterium. Deuterium is a hydrogen atom with a neutron. It has a molar mass twice that of hydrogen. Sketch the percent composition pie chart for heavy water.

SAMPLE PROBLEM

Calculating Percent Composition From a Chemical Formula

Propane (C_3H_8), the fuel commonly used in gas grills, is one of the compounds obtained from petroleum. Calculate the percent composition of propane.

ANALYZE List the knowns and the unknowns.

Knowns	Unknowns
mass of C in 1 mol C_3H_8 = 3 mol × 12.0 g/mol = 36.0 g	percent by mass of C = ?% C
mass of H in 1 mol C_3H_8 = 8 mol × 1.0 g/mol = 8.0 g	percent by mass of H = ?% H
molar mass of C_3H_8 = 36.0 g/mol + 8.0 g/mol = 44.0 g/mol	

CALCULATE Solve for the unknown.

Determine the percent by mass of C in the C_3H_8 by dividing the mass of C in 1 mole of the compound by the molar mass of the compound and multiplying by 100%.

$$\%C = \frac{\text{mass of C in 1 mol } C_3H_8}{\text{molar mass of } C_3H_8} \times 100\%$$

$$= \frac{36.0\ \cancel{g}}{44.0\ \cancel{g}} \times 100\% = 81.8\%\ C$$

Determine the percent by mass of H in C_3H_8.

$$\%H = \frac{\text{mass of H in 1 mol } C_3H_8}{\text{molar mass of } C_3H_8} \times 100\%$$

$$= \frac{8.0\ \cancel{g}}{44.0\ \cancel{g}} \times 100\% = 18\%\ H$$

EVALUATE Does the result make sense?

The percents of the elements add up to 100% when the answers are expressed to two significant figures (82% + 18% = 100%).

23 **SEP Use Mathematics** What is the percent by mass of nitrogen in the following fertilizers?

a. NH_3

b. NH_4NO_3

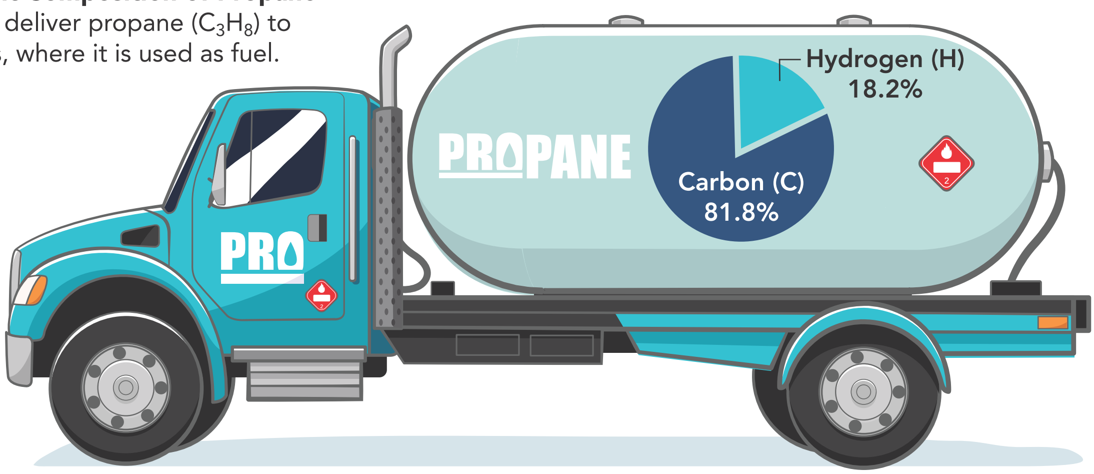
GO ONLINE for more practice problems.

Percent Composition as a Conversion Factor

You can use percent composition to calculate the number of grams of any element in a specific mass of a compound. The percent composition can be used as a conversion factor. For example, if you wanted to know the mass of hydrogen in 20 g of water, you would use the percent composition of hydrogen as a conversion factor. Because water is 11.1% hydrogen by mass, there are 11.1 g of hydrogen in 100 g of water.

$$20\ \cancel{g\ H_2O} \times \frac{11.1\ g\ H}{100\ \cancel{g\ H_2O}} = 2.2\ g\ H$$

Percent Composition of Propane
Trucks deliver propane (C_3H_8) to homes, where it is used as fuel.



In a 100-g sample of propane, you would have 81.8 g of carbon and 18.2 g of hydrogen.

$$\frac{81.8\ g\ C}{100\ g\ C_3H_8} \quad \text{and} \quad \frac{18.2\ g\ H}{100\ g\ C_3H_8}$$

24 **SEP Plan an Investigation** **Plan an investigation to determine the percent by mass of water that makes up a sample of hydrated copper(II) sulfate ($CuSO_4 \cdot 5H_2O$). Identify the steps you would take in your investigation.**

SAMPLE PROBLEM

Calculating the Mass of an Element in a Compound Using Percent Composition

Calculate the mass of carbon and the mass of hydrogen in 82.0 g of propane (C_3H_8).

ANALYZE List the known and the unknowns.

Knowns	Unknowns
mass of C_3H_8 = 82.0 g	mass of carbon = ? g C
	mass of hydrogen = ? g H

CALCULATE Solve for the unknowns.

Write the conversion factor to convert from mass of C_3H_8 to mass of C. The percent by mass of C in C_3H_8 is 81.8%.

$$\frac{81.8\ \text{g C}}{100\ \text{g } C_3H_8}$$

Multiply the mass of C_3H_8 by the conversion factor.

$$82.0\ \cancel{\text{g } C_3H_8} \times \frac{81.8\ \text{g C}}{100\ \cancel{\text{g } C_3H_8}} = \boxed{67.1\ \text{g C}}$$

Write the conversion factor to convert from mass of C_3H_8 to mass of H. The percent by mass of H in C_3H_8 is 18.2%.

$$\frac{18.2\ \text{g H}}{100\ \text{g } C_3H_8}$$

Multiply the mass of C_3H_8 by the conversion factor.

$$82.0\ \cancel{\text{g } C_3H_8} \times \frac{18.2\ \text{g H}}{100\ \cancel{\text{g } C_3H_8}} = \boxed{14.9\ \text{g H}}$$

EVALUATE Does the result make sense?

The sum of the two masses equals 82.0 g, the sample size.

25 **SEP Use Mathematics** **Calculate the grams of nitrogen in 125 g of each fertilizer.**

a. NH_3

b. NH_4NO_3

GO ONLINE for more practice problems.

Empirical Formulas

The **empirical formula** of a compound gives the lowest whole-number ratio of the atoms or moles of elements in a compound. For example, there are two atoms each of hydrogen and oxygen in one molecule of hydrogen peroxide (H_2O_2). The lowest ratio of hydrogen to oxygen is 1:1, so the empirical formula is HO.

An empirical formula may or may not be the same as the molecular formula, which gives the actual number of atoms in a molecule.

Empirical Formula From Data If you know the percent composition or the masses of the elements in a sample of an unknown compound, such as hydrogen peroxide, then you can determine the empirical formula. For example, suppose a sample of hydrogen peroxide contains 6% hydrogen and 94% oxygen.

The empirical formula of hydrogen peroxide shows that for every mole of hydrogen, there is 1 mole of oxygen.

Percent of element A		Mass of element A		Moles of element A		Mole ratio A-Z	Empirical formula
6% H = 0.06 H	$\times$ 100 g	6.0 g H	$\times \frac{1\ \text{mol H}}{1.0\ \text{g H}}$	6 mol H	$\div$ 5.88	1.02 H:1 O	HO
Percent of element Z		**Mass of element Z**		**Moles of element Z**			
94% O = 0.94 O	$\times$ 100 g	94 g O	$\times \frac{1\ \text{mol O}}{16.0\ \text{g O}}$	5.88 mol O	$\div$ 5.88		

Assume you have a 100-g sample. Multiply the decimal form of each percentage by 100 g to **find the mass of each element**.

Convert mass to moles by multiplying the mass of each element by its molar mass conversion factor.

Divide by the lowest number of moles to **determine the mole ratio**.

Write the empirical formula by converting the subscripts into whole numbers, if needed.

26 **CCC Structure and Function** Acetylene (C_2H_2) is a flammable gas used in welders' torches. Styrene (C_8H_8) is used to make packing peanuts. What is the empirical formula for each? Describe why the empirical formula might be useful in the lab setting but not useful for predicting the properties and/or functions of materials.

SAMPLE PROBLEM

Determining the Empirical Formula

An unknown compound is analyzed and found to contain 25.9% nitrogen and 74.1% oxygen. What is the empirical formula of the compound?

ANALYZE **List the knowns and the unknown.**

Knowns	Unknown
percent by mass of N = 25.9% N	empirical formula = $N_?O_?$
percent by mass of O = 74.1% O	

CALCULATE **Solve for the unknown.**

Convert the percent by mass of each element to moles to determine the mole ratio.

$$25.9\ \cancel{g\ N} \times \frac{1\ mol\ N}{14.0\ \cancel{g\ N}} = 1.85\ mol\ N$$

$$74.1\ \cancel{g\ O} \times \frac{1\ mol\ O}{16.0\ \cancel{g\ O}} = 4.63\ mol\ O$$

The mole ratio is $N_{1.85}$:$O_{4.63}$.

Divide each molar quantity by the smaller number of moles to get 1 mol from the element with the smaller number of moles.

$$\frac{1.85\ mol\ N}{1.85} = 1\ mol\ N$$

$$\frac{4.63\ mol\ O}{1.85} = 2.5\ mol\ O$$

The mole ratio becomes N_1:$O_{2.5}$.

Reduce the mole ratio to the lowest whole-number ratio by multiplying each part of the ratio by the smallest whole number that will convert both subscripts to a whole number.

1 mol N × 2 = 2 mol N

2.5 mol O × 2 = 5 mol N

The empirical formula is N_2O_5.

EVALUATE **Does the result make sense?**

The subscripts are whole numbers, and the percent composition of this empirical formula equals the percentages given in the original problem.

27 **SEP Use Mathematics** **Calculate the empirical formula for a compound that contains 67.6% Hg, 10.8% S, and 21.6% O.**

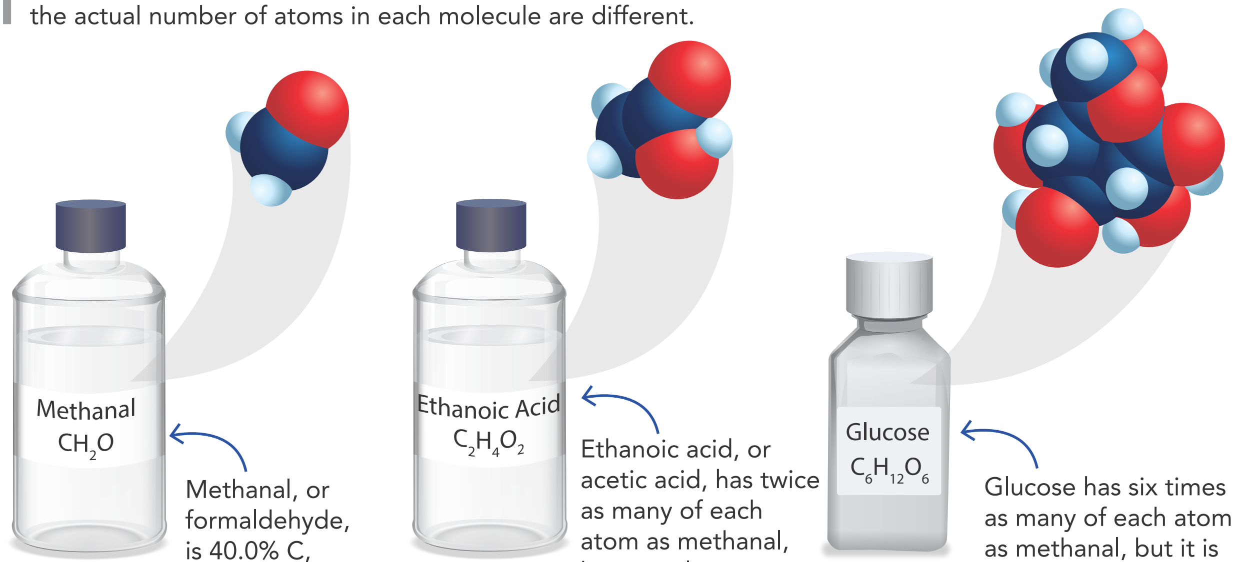
GO ONLINE for more practice problems.

Molecular Formulas

The **molecular formula** of a compound is either the same as its experimentally-determined empirical formula, or it is a simple whole-number multiple of its empirical formula. The molecular formula gives the actual number of atoms in the structure of the molecule, as opposed to just the ratio. For example, methanal, ethanoic acid, and glucose have the same empirical formula—CH_2O. However, they have different molecular formulas and very different properties.

You can determine the molecular formula of a compound from its empirical formula and its experimentally determined molar mass. Calculate the empirical formula mass (EFM) of a compound from its empirical formula. Then divide the molar mass by the empirical formula mass to find the number of empirical formula units in a molecule of the compound.

Empirical vs. Molecular Formula The empirical formula of methanal, ethanoic acid, and glucose shows that they have the same percent composition of C, H, and O. However, their molecular formulas show that the actual number of atoms in each molecule are different.



Methanal, or formaldehyde, is 40.0% C, 6.7% H, and 53.3% O.

Ethanoic acid, or acetic acid, has twice as many of each atom as methanal, but it is also 40.0% C, 6.7% H, and 53.3% O.

Glucose has six times as many of each atom as methanal, but it is also 40.0% C, 6.7% H, and 53.3% O.

28 **SEP Interpret Data** **If you know the molar mass of methanal, how could you determine the molar mass of ethanoic acid and glucose? Complete the table.**

Formula (name)	Molar mass (g/mol)
CH	13
C_2H_2 (ethyne)	26 (2 × 13)
C_6H_6 (benzene)	78 (6 × 13)
CH_2O (methanal)	30
$C_2H_4O_2$ (ethanoic acid)	
$C_6H_{12}O_6$ (glucose)	

SAMPLE PROBLEM

Determining the Molecular Formula

Calculate the molecular formula of a compound whose molar mass is 60.0 g/mol and empirical formula is CH_4N.

ANALYZE List the knowns and the unknown.

Knowns	Unknown
empirical formula = CH_4N	molecular formula = $C_?H_?N_?$
molar mass = 60.0 g/mol	

CALCULATE Solve for the unknown.

Calculate the empirical formula mass (EFM). It is the molar mass of the empirical formula.

EFM of CH_4N =
12.0 g/mol + 4(1.0 g/mol) + 14.0 g/mol = 30.0 g/mol

Divide the molar mass by the empirical formula mass to obtain a whole number.

$$\frac{\text{molar mass}}{\text{EFM}} = \frac{60.0\ \cancel{\text{g/mol}}}{30.0\ \cancel{\text{g/mol}}} = 2$$

Multiply the empirical formula subscripts by this value to get the molecular formula.

$(CH_4N) \times 2 = C_2H_8N_2$

EVALUATE Does the result make sense?

The molecular formula has the molar mass of the compound.

29 **SEP Use Mathematics** Find the molecular formula of ethylene glycol, which is used in antifreeze. The molar mass is 62.0 g/mol, and the empirical formula is CH_3O.

30 **SEP Use Mathematics** What is the molecular formula of a compound with the empirical formula CClN and a molar mass of 184.5 g/mol?

GO ONLINE for more practice problems.

Revisit

INVESTIGATIVE PHENOMENON

GO ONLINE to Elaborate on and Evaluate your knowledge of percent composition and the conservation of mass by completing the peer review and engineering design activities.

In the CER worksheet you completed at the beginning of the investigation, you drafted an explanation for the ways in which matter is quantified. With a partner, reevaluate your arguments.

31 **SEP Obtain and Evaluate Information** Monosaccharides are the basic units of carbohydrates called "simple sugars." Two such monosaccharides are glucose, found in plants and animals, and fructose, found in fruits and honey. Both glucose and fructose have the same empirical formula and the same molecular formula ($C_6H_{12}O_6$). However, they have different structures, which is why they have different names. What are the empirical formulas for glucose and fructose? Research and compare their structures and list two other carbohydrates that have the chemical formula $C_6H_{12}O_6$. How do their structures affect their functions?

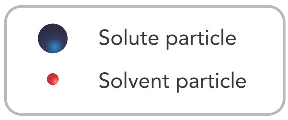
Concentrations of Solutions

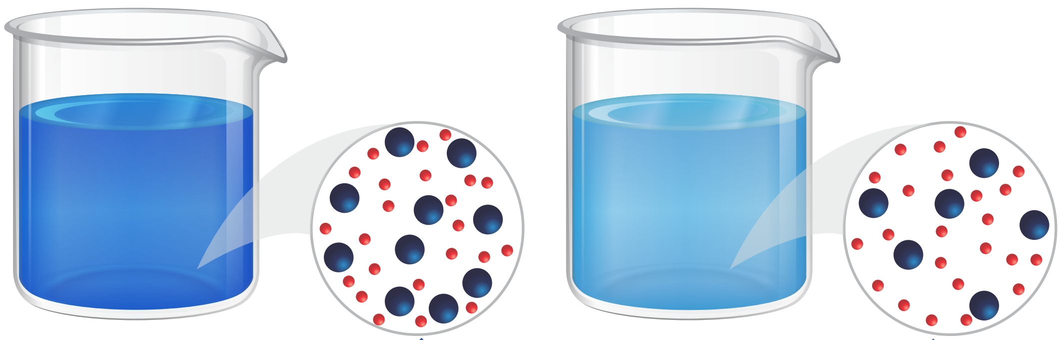
GO ONLINE to Explore and Explain concentration and molarity.

Molarity

Concentration of a Solution The **concentration** of a solution is a measure of the amount of solute that is dissolved in a given quantity of solvent. You can describe the amount of a solute in solution qualitatively using the terms *dilute* and *concentrated*. A solution with a small amount of solute is a **dilute solution.** By contrast, a **concentrated solution** contains a large amount of solute. A solution with 5 g copper sulfate per 100 mL water might be described as dilute compared to a solution with 10 g copper sulfate per 100 mL water.

Amount of Solute The more solute particles in a given amount of solution, the greater the concentration of the solute.

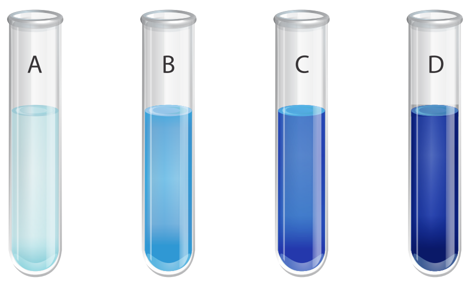




In a **concentrated** solution, there are many solute particles in the solvent.

In a **dilute** solution, there are few solute particles in the solvent.

32 **SEP Define a Problem** The test tubes contain water (A) and three copper sulfate solutions (B–D) with increasing concentration. Why is it difficult to say whether solution C is a concentrated solution or a dilute solution? How can we better define the problem?



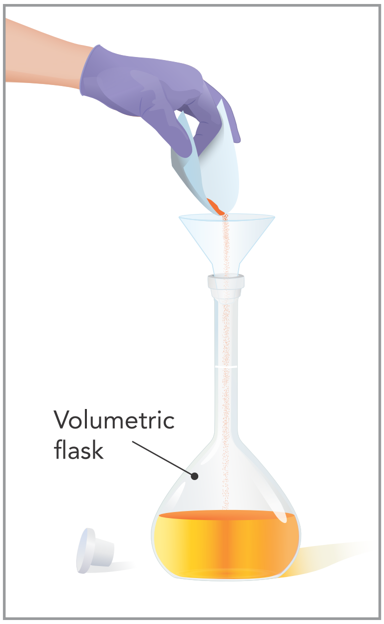
Defining Concentration Quantitatively Molar concentration, called **molarity** (M), is the number of moles of solute dissolved in 1 liter of solution. To calculate the molarity of a solution, divide the number of moles of solute by the volume of the solution in liters.

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

A common error is to use the volume of the solvent in the denominator. The volume of the solution contains both the solute and solvent.

Making a 0.5 *M* Solution Volumetric flasks are used to make specific volumes of solution because they hold a precise amount of liquid. A line on the neck of the flask marks the volume.

Step 1 Start by filling a 1 L volumetric flask halfway with distilled water. Then add 0.5 mol of solute.



Step 2 Mix the solute and solvent by gently swirling the flask to dissolve the solid solute.

The solute needs to be completely dissolved to ensure the correct final volume of solution.

Step 3 Carefully fill the flask with distilled water to the 1 L mark. Note that you have made 1 L of solution, not added 1 L of solvent.

33 **SEP Carrying out Investigations** **A student adds 1 L of water to 1 mol of sodium chloride. Will the result be a 1 *M* solution of sodium chloride and water? Why or why not?**

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SAMPLE PROBLEM

Calculating Molarity

Intravenous saline solutions are often administered to patients in the hospital. One saline solution contains 0.90 g NaCl in exactly 100 mL of solution. What is the molarity of the solution?

ANALYZE List the knowns and the unknown.

Knowns	Unknown
solution concentration = 0.90 g NaCl/100 mL	solution concentration = ? *M*
molar mass of NaCl = 58.5 g/mol	

CALCULATE Solve for the unknown.

Identify the steps to convert concentration from g/100 mL to mol/L.

g/100 mL → mol/100 mL → mol/L

Use the molar mass to convert g NaCl/100 mL to mol NaCl/100 mL. Convert the volume units from mL to L.

solution concentration =

$$\frac{0.90\ \cancel{\text{g NaCl}}}{100\ \cancel{\text{mL}}} \times \frac{1\text{ mol NaCl}}{58.5\ \cancel{\text{g NaCl}}} \times \frac{1000\ \cancel{\text{mL}}}{1\text{ L}} =$$

$$0.15\ \frac{\text{mol}}{\text{L}} = \boxed{0.15\ M}$$

The relationship 1 L = 1000 mL gives you the conversion factor 1000 mL/1 L.

EVALUATE Does the result make sense?

The answer should be less than 1 *M* because 0.90 g/100 mL is the same concentration as 9.0 g/L, and 9.0 g is less than 1 mol of NaCl. The answer is correctly expressed to two significant figures.

34 **SEP Use Mathematics** A solution has a volume of 2.0 L and contains 36.0 g of glucose ($C_6H_{12}O_6$). If the molar mass of glucose is 180 g/mol, what is the molarity of the solution?

35 **SEP Use Mathematics** A solution has a volume of 250 mL and contains 0.70 mol NaCl. What is its molarity?

GO ONLINE for more practice problems.

SAMPLE PROBLEM

Calculating Moles of Solute in Solution

Household laundry bleach is a dilute aqueous solution of sodium hypochlorite ($NaClO$). How many moles of solute are present in 1.5 L of 0.70 *M* $NaClO$?

ANALYZE List the knowns and the unknown.

Knowns	Unknown
volume of solution = 1.5 L	moles solute = ? mol
solution concentration = 0.70 *M* $NaClO$	

CALCULATE Solve for the unknown.

Identify the steps to convert volume to moles using molarity as a conversion factor.

volume of solution → moles of solute

Multiply the given volume by the molarity expressed in mol/L.

$$\text{moles solute} = 1.5\ \cancel{\text{L}} \times \frac{0.70\ \text{mol NaClO}}{1\ \cancel{\text{L}}} = 1.1\ \text{mol NaClO}$$

EVALUATE Does the result make sense?

The answer should be greater than 1 mol but less than 1.5 mol because the solution concentration is less than 0.75 mol/L and the volume is less than 2 L. The answer is correctly expressed to two significant figures.

36 **SEP Use Mathematics** How many moles of ammonium nitrate are in 335 mL of 0.425 *M* NH_4NO_3?

37 **SEP Use Mathematics** How many moles of solute are in 250 mL of 2.0 *M* $CaCl_2$?

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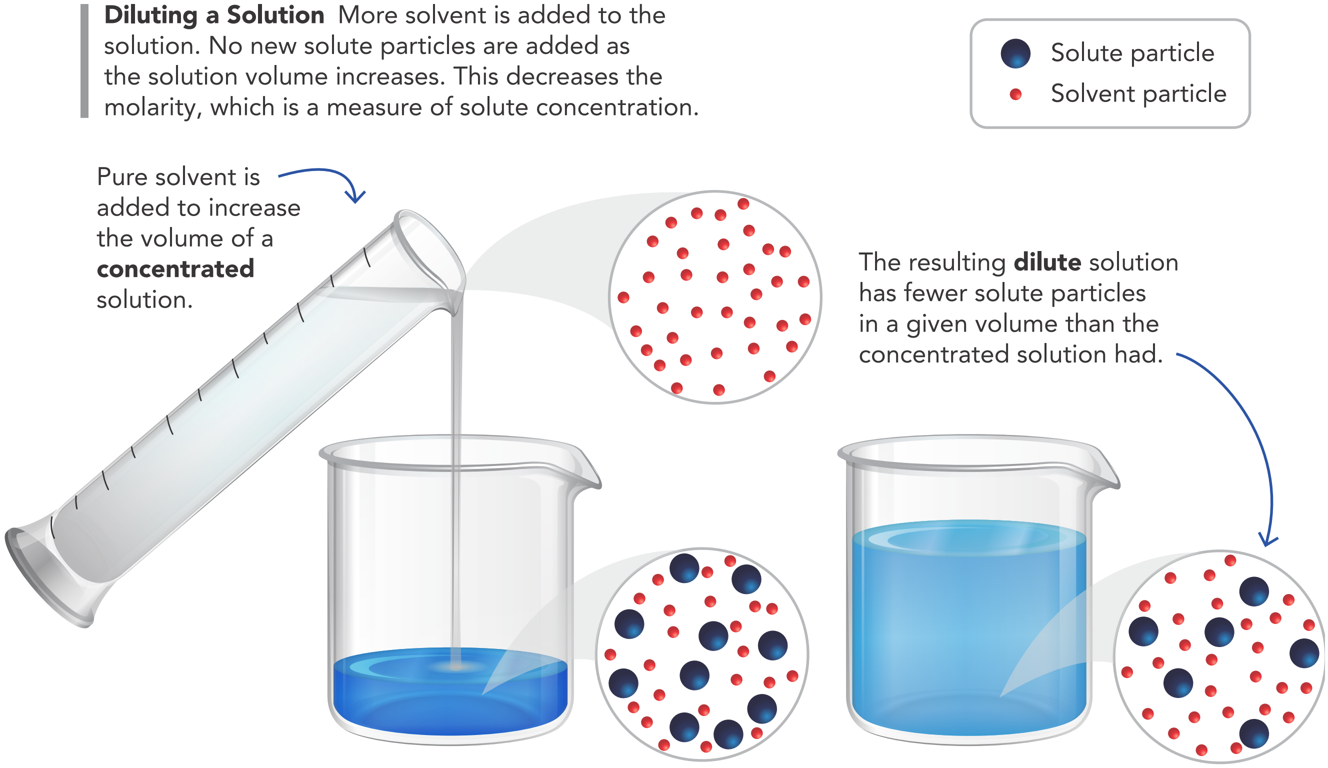
Dilutions

Diluting a Solution You dilute a solution by adding more solvent to it. During dilution, the total number of moles of solute does not change. However, the number of moles of the solvent does change.

Therefore, diluting a solution reduces the number of moles of solute per unit volume but not the total number of moles of solute.

Diluting a Solution More solvent is added to the solution. No new solute particles are added as the solution volume increases. This decreases the molarity, which is a measure of solute concentration.



38 **SEP Develop a Model** Imagine you have a box and a set of red and green balls. The red balls represent solute particles, and the green balls represent solvent particles. Sketch a model showing how the number of red balls doesn't change as more green balls are added. Use your model to show how the percent by count of red balls to total balls decreases.

Making a Dilution The number of moles of solute is related to the molarity (M) and volume (V) of a solution by the following equation:

$$\text{moles of solute} = M \times V$$

A stock solution is a concentrated solution of known molarity that is diluted to make less concentrated solutions. Because the molarity is known, a defined volume of the stock can be measured to make a dilution with the desired molarity. Since the total number of moles of solute does not change during a dilution, we can write the following expression:

$$\text{moles of solute} = \underbrace{M_{stock}}_{\text{Molarity of the stock solution}} \times \underbrace{V_{stock}}_{\text{Volume of the stock solution}} = \underbrace{M_{dilute}}_{\text{Molarity of the dilute solution}} \times \underbrace{V_{dilute}}_{\text{Volume of the dilute solution}}$$

Making a Dilution Dilutions often require making a solution with a precise concentration from a small amount of stock solution.

Step 1 A volume (V_{stock}) containing the desired moles of solute is measured from a stock solution of known concentration (M_{stock}).

Step 2 The measured volume is transferred to another volumetric flask that will measure the desired amount of dilute solution.

Step 3 Solvent is carefully added to the mark on the flask to make a dilute solution with the desired concentration (M_{dilute}). The volume of solution in the flask is V_{dilute}.

39 **SEP Interpret Data** The table shows the volume (V_{stock}) of a 2.0 M $MgSO_4$ stock solution needed to make several different 1 L dilutions at the molarities (M_{dilute}) shown. Complete the table by determining how much stock solution you should measure out to make 1 L of Dilution 3.

Dilutions From a 2.0 M Stock Solution

Dilution	V_{stock} (L)	$M \times V$ (mol)	M_{dilute} (mol/L)	V_{dilute} (L)
Dilution 1	0.5	1	1	1
Dilution 2	0.375	0.75	0.75	1
Dilution 3			0.5	1

SAMPLE PROBLEM

Preparing a Dilute Solution

You have a stock solution of aqueous 2.00 *M* $MgSO_4$. How many milliliters must be diluted with water to prepare 100.0 mL of aqueous 0.400 *M* $MgSO_4$?

ANALYZE List the knowns and the unknown.

Knowns	Unknown
$M_{stock} \times V_{stock} = M_{dilute} \times V_{dilute}$	V_{stock} = ? mL of 2.00 *M* $MgSO_4$
M_{stock} = 2.00 *M* $MgSO_4$	
M_{dilute} = 0.400 *M* $MgSO_4$	
V_{dilute} = 100.0 mL of 0.400 *M* $MgSO_4$	

CALCULATE Solve for the unknown.

Rearrange the equation to solve for V_{stock}.

$$V_{stock} = \frac{M_{dilute} \times V_{dilute}}{M_{stock}}$$

Substitute known values into the equation.

$$V_{stock} = \frac{0.400\ \cancel{M} \times 100.0\ \text{mL}}{2.00\ \cancel{M}} = 20.0\ \text{mL}$$

EVALUATE Does the result make sense?

The initial concentration is five times larger than the dilute concentration. Because the number of moles of solute does not change, the initial volume of solution should be one fifth the final volume of the diluted solution.

40 **SEP Design a Solution** How could you prepare 250 mL of 0.20 *M* NaCl using only a solution of 1.0 *M* NaCl and water?

GO ONLINE for more practice problems.

Percent Solution

If both the solute and solvent are liquids, a convenient way to make a solution is to measure the volumes of the solute and solution. **Percent by volume** of a solution is the ratio of the volumes of solute to solution.

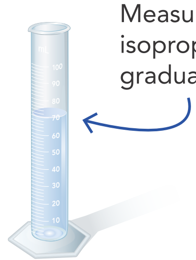
$$\text{Percent by volume } [\%\ (v/v)] = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$$

The **percent by mass** of a solution is the ratio of the masses of the solute and solution.

$$\text{Percent by mass } [\%\ (m/m)] = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

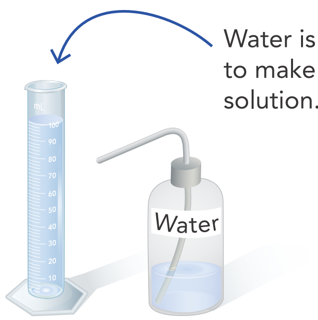
Percent by Volume Isopropyl alcohol is sold as a 70% solution by volume. The concentration is written as 70% by volume or 70% (v/v).

Measure 70 mL isopropyl alcohol in a graduated cylinder.



Pure isopropyl alcohol

Water is added to make a 100 mL solution.

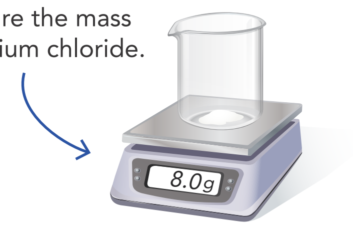


The percent by volume of the solution is

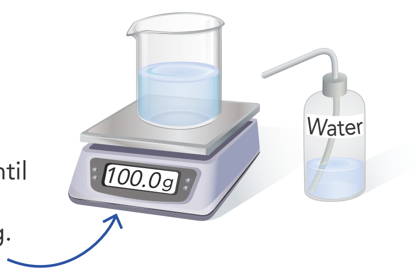
$$\frac{\text{70 mL isopropyl alcohol}}{\text{100 mL solution}} \times 100 = 70\% \text{ isopropyl alcohol (v/v).}$$

Percent by Mass A brine solution consists of 8.0 g NaCl in 100 g of solution and has 8.0% sodium chloride by mass. The concentration is written as 8.0% by mass or 8.0% (m/m).

Measure the mass of sodium chloride.



Water is added until the mass of the solution is 100.0 g.



The percent by mass of the solution is

$$\frac{\text{8.0 g sodium chloride}}{\text{100 g solution}} \times 100 = 8.0\% \text{ NaCl (m/m).}$$

41 **SEP Design a Solution** Professional car detailers use a 50% (v/v) solution of isopropyl rubbing alcohol to remove old wax from the car surface. Design a procedure to make a 50% (v/v) solution from store-bought rubbing alcohol, which is 70% (v/v), that fills a 1-L spray bottle.

SAMPLE PROBLEM

Calculating Percent by Volume

What is the percent by volume of ethanol (C_2H_6O, or ethyl alcohol) in the final solution when 85 mL of ethanol is diluted to a volume of 250 mL with water?

ANALYZE List the knowns and the unknown.

Knowns	Unknown
volume of solute = 85 mL ethanol	percent by volume = ?% ethanol (v/v)
volume of solution = 250 mL ethanol	

CALCULATE Solve for the unknown.

Write the equation for percent by volume.

$$\text{Percent by volume [\% (v/v)]} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$$

Substitute known values into the equation.

$$\%\ (\text{v/v}) = \frac{85\ \text{mL ethanol}}{250\ \text{mL}} \times 100\% = 34\%\ \text{ethanol (v/v)}$$

EVALUATE Does the result make sense?

The volume of the solute is about one third the volume of the solution, so the answer is reasonable. The answer is correctly expressed to two significant figures.

42 **SEP Design a Solution** A bottle of hydrogen peroxide (H_2O_2) is labeled 3.0% (v/v). How many mL H_2O_2 are in a 400.0-mL bottle?

GO ONLINE for more practice problems.

Revisit

INVESTIGATIVE PHENOMENON

GO ONLINE to Elaborate on and Evaluate your knowledge of solution concentration by completing the discussion and data analysis activities.

In the CER worksheet you completed at the beginning of the investigation, you drafted an explanation for the ways in which matter is quantified. With a partner, reevaluate your arguments.

43 **SEP Plan an Investigation** You make a pitcher of 2 L lemonade following a recipe, but you find that it tastes too sweet. Plan an investigation where you dilute the lemonade into different dilutions with varying concentrations of sugar. Once you have determined your ideal dilution, describe how and why you would adjust the recipe or use the pitcher of lemonade to make a glass of lemonade in the future.

GO ONLINE to Evaluate what you learned about using mathematical representations to quantify matter by using the available assessment resources.

In the Performance-Based Assessment, you collected evidence for determining chemical formulas based on percent composition. Wrap up your analysis by answering the following questions.

44 **SEP Identify Limitations of Models** Reevaluate your evidence. How can you improve the methodology of collecting the data and the quality of the data collected?

45 **Revisit the Anchoring Phenomenon** How does what you learned in this investigation help you understand how the quantification of matter can be used to design better foods?