The Mole

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CHAPTER 1 The Mole

CHAPTER OUTLINE

Look at the box of fresh donuts in the picture. If you have a sweet tooth, they probably look delicious. When you go to a donut shop or to the grocery store, you typically buy donuts according to the number—a single donut, by the half-dozen, or by the dozen. That makes it easy for both the seller and the buyer. However, it would be possible to price the donuts and sell them according to mass. A jelly-filled donut would probably weigh more and thus be more expensive than a sugared or a powdered donut. Donut shops don't do this because it would take too much time. In this chapter you will learn about different ways to measure the amount of something and the interrelationships between those measurements. You will also be introduced to one of the most important units used in chemistry, the mole.

1.1 The Mole Concept

Lesson Objectives

- Identify three methods for measuring the amount of matter in a sample.
- Define the mole and its relationship to Avogadro's number.
- Use Avogadro's number to convert between moles and the number of representative particles of a substance.
- Relate the atomic mass of an element to its molar mass.
- Calculate the molar mass of a given compound.

Lesson Vocabulary

- Avogadro's number
- formula mass
- molar mass
- mole
- representative particle

Check Your Understanding

Recalling Prior Knowledge

- What is a conversion factor? What is dimensional analysis?
- What is meant by the atomic mass of an element, and in what units are atomic masses expressed?
- How are the structures of molecular compounds and ionic compounds different?

Chemistry is a quantitative science. It is not enough to simply observe chemical reactions and describe what happens. Chemists always need to know how much. How many liters of carbon dioxide gas are going to be produced for every gallon of gasoline that is burned? How many kilograms of the industrial chemical sulfuric acid are produced in a typical year? How many kilograms of sulfur, oxygen, and water are required to manufacture that much sulfuric acid? These types of questions show the quantitative nature of chemistry and chemical reactions.

How Much Matter?

If you do any baking, you may have a set of canisters in your kitchen that hold large amounts of flour and sugar so that they are easy to access when preparing some cookies or a cake. Think about how much sugar is in one of those canisters. How would you measure it? There are actually multiple answers to this question. One way to measure the amount of sugar in the canister is to find its mass. Another would be to measure its volume. A third way, and a very time consuming one, would be to count all of the individual grains of sugar. The amount of matter can be measured in three basic ways: mass, volume, and number of particles. In order to fully understand and manipulate chemical reactions, chemists must be able to understand these three ways of measuring matter and the interconnections between them.

Names for Numbers

How many bananas is a dozen bananas (**Figure** 1.1)? How many elephants is a dozen elephants? How many asteroids is a dozen asteroids? For all of these questions, the answer is obviously twelve. There is nothing that we need to know about the bananas or the elephants or the asteroids in order to answer the question. The term dozen is always used to refer to twelve of something; it is a name that is given to an amount. Other amounts are given special names as well. For example, a pair is always two, and a gross of something is a dozen dozens, which would be 144.

FIGURE 1.1
Bananas can be sold by mass or by count.

We can use a conversion factor and dimensional analysis to convert back and forth between the number of items and the name given to a certain number. For example, if you wanted to know how many bananas there are in 8 dozens, you could perform the following calculation:

$$8 \text{ dozen bananas} \times \frac{12 \text{ bananas}}{1 \text{ dozen bananas}} = 96 \text{ bananas}$$

The conversion factor of 12 items = 1 dozen items is true regardless of the identity of the item. Alternatively, you could find out how many dozens of asteroids are in a collection of 1242 asteroids.

$$1242 \text{ asteroids} \times \frac{1 \text{ dozen asteroids}}{12 \text{ asteroids}} = 103.5 \text{ dozen asteroids}$$

The conversion factor is simply inverted in this case so that the units cancel correctly.

Conversion factors can also be used to relate the amount of something to its mass. Suppose that you have a small bunch of bananas consisting of five bananas. You place them on a balance and find that the five bananas have a mass of 850 g. Assuming each banana has the same mass, what would be the mass of 3 dozen bananas? We can employ two conversion factors to find the answer.

$$3 \text{ dozen bananas} \times \frac{12 \text{ bananas}}{1 \text{ dozen bananas}} \times \frac{850 \text{ g}}{5 \text{ bananas}} = 6120 \text{ g} = 6100 \text{ g}$$

By knowing the mass of five bananas, we now have a relationship that we can use to convert between mass and number of bananas for any number or any mass. Note that because 850 g was a measured quantity with two significant figures, the result was also rounded to two significant figures. The amounts 5 and 12 are found by counting, so they are exact quantities and have unlimited significant figures.

The Mole and Avogadro's Number

It certainly is easy to count bananas or to count elephants (as long as you stay out of their way). However, you would be counting grains of sugar from your sugar canister for a long, long time. Recall from the chapter, *Atomic Structure* that atoms and molecules are extremely small—far, far smaller than grains of sugar. Counting atoms or molecules is not only unwise, it is absolutely impossible. One drop of water contains about 10^{22} molecules of water. If you counted 10 molecules every second for 50 years without stopping, you would have counted only 1.6×10^{10} molecules. At that rate, it would take you over 30 trillion years to count the water molecules in one tiny drop!

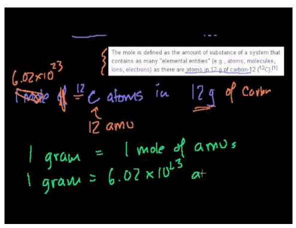
Chemists needed a name that can stand for a very large number of items. Amedeo Avogadro (1776-1856), an Italian scientist (**Figure** 1.2), provided just such a number. He is responsible for the counting unit of measure called the mole (**Figure** 1.3). A **mole** (mol) *is the amount of a substance that contains* 6.02×10^{23} *representative particles of that substance. The mole is the SI unit for amount of a substance*. Just like the dozen and the gross, it is a name that stands for a number. There are therefore 6.02×10^{23} water molecules in a mole of water molecules. There also would be 6.02×10^{23} bananas in a mole of bananas, if such a huge number of bananas ever existed.

FIGURE 1.2
Italian scientist Amedeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry.

The number 6.02×10^{23} is called **Avogadro's number**, *the number of representative particles in a mole*. It is an experimentally determined number. A **representative particle** *is the smallest unit in which a substance naturally exists*. For the majority of pure elements, the representative particle is the atom. Samples of pure iron, carbon, and helium consist of individual iron atoms, carbon atoms, and helium atoms, respectively. Seven elements exist in nature as diatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2), so the representative particle for these elements is the molecule. Likewise, all molecular compounds, such as H_2O and CO_2, exist as molecules, so the molecule is their representative particle as well. For ionic compounds such as NaCl and $Ca(NO_3)_2$, the representative particle is the formula unit. A mole of any substance contains Avogadro's number (6.02×10^{23}) of representative particles.

You can watch a video lecture about moles (the unit) by going to http://www.youtube.com/watch?v=AsqEkF7hcII
.

FIGURE 1.3
The animal mole (*left*) is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot (*right*). National Mole Day is a celebration of chemistry that occurs on October 23rd (10/23) of each year.



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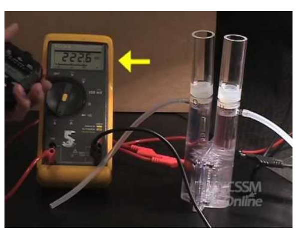
Watch a video that shows an experimental calculation of Avogadro's Number: http://www.youtube.com/watch?vp9QYJqFq5s .



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Video experiment questions:

1. Briefly describe this experiment.
2. Write the calculations shown in this experiment.
3. What was the experimental result?
4. What was the percent error?

Conversions Between Moles and Number of Particles

Just as we did with dozens of bananas, we can use the number of items in a mole to convert back and forth between a number of particles and moles of those particles.

Sample Problem 10.1: Converting a Number of Particles to Moles

The element, carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms are in a sample containing 4.72×10^{24} atoms of carbon?

Step 1: List the known quantities and plan the problem.

Known

- number of C atoms = 4.72×10^{24}

- 1 mole = 6.02×10^{23}

Unknown

- 4.72×10^{24} C atoms = ? mol C

One conversion factor will allow us to convert from the number of C atoms to moles of C atoms.

Step 2: Calculate.

$$4.72 \times 10^{24} \text{ atoms C} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ atoms C}} = 7.84 \text{ mol C}$$

Step 3: Think about your result.

The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is greater than 1 mole. Since our starting value is reported to three significant figures, the result of the calculation is also rounded to three significant figures.

Practice Problems

1. Convert the given number of particles to moles.
   a. 3.65×10^{22} molecules of H_2O
   b. 9.18×10^{23} formula units of KCl
2. Convert the given number of moles to the number of representative particles.
   a. 1.25 mol Zn atoms
   b. 0.061 mol CH_4 molecules

Suppose that you wanted to know how many hydrogen atoms were in a mole of water molecules. First, you would need to know the chemical formula for water, which is H_2O. There are two atoms of hydrogen in each molecule of water. How many atoms of hydrogen would there be in two water molecules? There would be $2 \times 2 = 4$ hydrogen atoms (**Figure** 1.4). How about in a dozen? Since a dozen is 12, there would be $12 \times 2 = 24$ hydrogen atoms in a dozen water molecules. To get the answers, (4 and 24) you had to multiply the given number of molecules by two atoms of hydrogen per molecule. Similarly, to find the number of hydrogen atoms in a mole of water molecules, the problem could be solved using conversion factors.

$$1 \text{ mol } H_2O \times \frac{6.02 \times 10^{23} \text{ molecules } H_2O}{1 \text{ mol } H_2O} \times \frac{2 \text{ atoms H}}{1 \text{ molecule } H_2O} = 1.20 \times 10^{24} \text{ atoms H}$$

The first conversion factor converts from moles of particles to the number of particles. The second conversion factor reflects the number of atoms contained within each molecule.

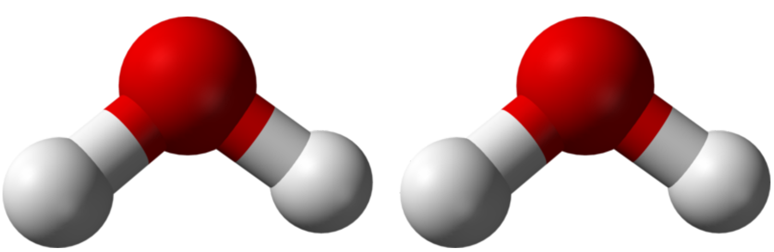


FIGURE 1.4

Two water molecules contain 4 hydrogen atoms and 2 oxygen atoms. A mole of water molecules contains 2 moles of hydrogen atoms and 1 mole of oxygen atoms.

Sample Problem 10.2: Atoms, Molecules, and Moles

Sulfuric acid has the chemical formula H_2SO_4. A certain quantity of sulfuric acid contains 4.89×10^{25} atoms of oxygen. How many moles of sulfuric acid are in the sample?

Step 1: List the known quantities and plan the problem.

Known

- number of O atoms in the sample = 4.89×10^{25}
- 1 mol H_2SO_4 = 6.02×10^{23} molecules H_2SO_4

Unknown

- mol of H_2SO_4 molecules in the sample

Two conversion factors will be used. First, convert atoms of oxygen to molecules of sulfuric acid. Then, convert molecules of sulfuric acid to moles of sulfuric acid.

Step 2: Calculate.

$$4.89 \times 10^{25} \text{ atoms O} \times \frac{1 \text{ molecule } H_2SO_4}{4 \text{ atoms O}} \times \frac{1 \text{ mol } H_2SO_4}{6.02 \times 10^{23} \text{ molecules } H_2SO_4} = 20.3 \text{ mol } H_2SO_4$$

Step 3: Think about your result.

The original number of oxygen atoms was about 80 times larger than Avogadro's number. Since each sulfuric acid molecule contains 4 oxygen atoms, there are about 20 moles of sulfuric acid molecules.

Practice Problems

3. How many atoms of carbon are in 0.750 moles of propane, which has a chemical formula of C_3H_8?
4. The chemical formula of glucose is $C_6H_{12}O_6$. How many moles of glucose are present in a sample that contains 2.46×10^{24} atoms of hydrogen?

Molar Mass

Because we are not able to count individual atoms, it is important to have a way to convert between amounts, which are expressed in moles, and a unit of quantity that we can more easily measure, such as mass. We begin by looking at the periodic table, which tells us the relative masses of various elements.

Molar Masses of Elements

As you learned previously, the atomic masses found on the periodic table are in atomic mass units. For example, one atom of the most abundant isotope of hydrogen has a mass of approximately 1 amu, and one atom of helium has a mass of about 4 amu. Atomic masses are relative masses; they are based on the definition that one amu is equal to $1/12^{th}$ of the mass of a single atom of carbon-12. Therefore, one atom of carbon-12 has a mass of 12 amu, which is three times heavier than an atom of helium. This ratio would hold for any number of carbon and helium atoms. One hundred carbon-12 atoms would have three times the mass of one hundred helium atoms. By extension, 12.00 g of carbon-12 would contain the same number of atoms as 4.00 g of helium.

The relative scale of atomic masses in amu is also a relative scale of masses in grams. An alternative definition of the mole is that it is the amount of a substance that contains as many representative particles as the number of atoms in exactly 12 g of carbon-12. In other words, exactly 12 g of carbon-12 contains one mole, or 6.02×10^{23} atoms of carbon-12. Likewise, 4.00 g of helium also contains one mole, or 6.02×10^{23} atoms of helium. *The atomic mass of an element, expressed in grams, is the mass of one mole of that element*. **Molar mass** *is defined as the mass of one mole of representative particles of a substance*. By looking at the periodic table, we can see that the molar mass of lithium is 6.94, the molar mass of zinc is 65.38, and the molar mass of gold is 196.97. Each of these quantities contains 6.02×10^{23} atoms of that particular element. The units for molar mass are grams per mole (g/mol). For example, the molar mass of zinc is 65.38 g/mol.

Recall that the atomic masses on the periodic table are generally not whole numbers because each atomic mass is a weighted average of all the naturally occurring isotopes of that element. Since any usable quantity of an element contains a very, very large number of atoms, those weighted averages in grams can be used as the molar mass of the element. For our purposes, we will use the molar masses rounded to the hundredths place (two digits after the decimal point).

Molar Masses of Compounds

The molecular formula of carbon dioxide is CO_2. One molecule of carbon dioxide consists of 1 atom of carbon and 2 atoms of oxygen. We can calculate the mass of one molecule of carbon dioxide by adding together the masses of 1 atom of carbon and 2 atoms of oxygen.

12.01 amu + 2(16.00 amu) = 44.01 amu

The molecular mass of a compound is the mass of one molecule of that compound. The molecular mass of carbon dioxide is 44.01 amu.

Recall that ionic compounds do not exist as discrete molecules, but rather as an extended three-dimensional network of ions called a crystal lattice. The empirical formula of an ionic compound tells us the ratio of the ions in the crystal. The mass of one formula unit of an ionic compound is called the formula mass. The formula mass of sodium sulfide, Na_2S, can be calculated as follows:

2(22.99 amu) + 32.06 amu = 78.04 amu

The **formula mass** *is the sum of the masses of all the atoms represented in a chemical formula*. The term formula mass is applicable to molecular compounds, ionic compounds, or ions.

The molar mass of any compound is the mass in grams of one mole of that compound. One mole of carbon dioxide molecules has a mass of 44.01 g, while one mole of sodium sulfide formula units has a mass of 78.04 g. Their molar masses are 44.01 g/mol and 78.04 g/mol, respectively. In both cases, that is the mass of 6.02×10^{23} representative particles. The representative particle of CO_2 is the molecule, while for Na_2S, it is the formula unit.

Sample Problem 10.3: Molar Mass of a Compound

Calcium nitrate, $Ca(NO_3)_2$, is used as a component in fertilizer. Determine the molar mass of calcium nitrate.

Step 1: List the known and unknown quantities and plan the problem.

Known

- formula = $Ca(NO_3)_2$
- molar mass of Ca = 40.08 g/mol
- molar mass of N = 14.01 g/mol
- molar mass of O = 16.00 g/mol

Unknown

- molar mass of $Ca(NO_3)_2$

First, we need to analyze the formula. Since Ca lacks a subscript, there is one Ca atom per formula unit. The 2 outside the parentheses means that there are two nitrate ions per formula unit, and each nitrate ion consists of one nitrogen atom and three oxygen atoms. Therefore, there are a total of $1 \times 2 = 2$ nitrogen atoms and $3 \times 2 = 6$ oxygen atoms per formula unit. Thus, 1 mol of calcium nitrate contains 1 mol of Ca atoms, 2 mol of N atoms, and 6 mol of O atoms.

Step 2: Calculate.

Use the molar mass of each atom together with the quantity of each atom in the formula to find the total molar mass.

$$\begin{aligned} 1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} &= 40.08 \text{ g Ca} \\ 2 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} &= 28.02 \text{ g N} \\ 6 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} &= 96.00 \text{ g O} \end{aligned}$$

molar mass of $Ca(NO_3)_2$ = 40.08 g + 28.02 g + 96.00 g = 164.10 g/mol

Step 3: Think about your result.

The molar mass is the mass in grams of 1 mol of calcium nitrate. It is expressed to the hundredths place because the numbers being added together are expressed to the hundredths place.

Practice Problem

5. Calculate the molar masses of the following compounds.
    1. C_2H_6
    2. $(NH_4)_2SO_4$

Lesson Summary

- The amount of matter in a given sample can be measured by its mass, volume, or the number of particles.
- A mole of any substance contains Avogadro's number (6.02×10^{23}) representative particles of the substance. A representative particle can be an atom, an ion, a molecule, or a formula unit.
- The molar mass of an element is its atomic mass expressed in grams and is equal to the mass of one mole of atoms of that element.
- The molar mass of a compound is the mass of one mole of representative particles of the compound. Molar mass is calculated by multiplying the molar mass of each element in the compound by the number of atoms of that element present in one formula unit and adding the resulting values together.

Lesson Review Questions

Reviewing Concepts

1. What is Avogadro's number and what does it represent?
2. What is the representative particle for each of the following substances?
   a. barium chloride
   b. silicon
   c. nitrogen gas
   d. water
3. How many oxygen atoms are there in a representative particle of each of the following substances?
   a. $KClO_4$
   b. CH_3COOH
   c. $Al(NO_3)_3$
4. What is wrong with the following statements?
   a. A mole of any substance contains the same number of atoms.
   b. One mole of water contains Avogadro's number of atoms.
5. If the atomic mass of a certain element is 69.72 amu, what is its molar mass?

Problems

6. Calculate the number of representative particles in each of the following.
   a. 0.0391 mol Ne
   b. 3.72 mol NH_3
   c. 8.00 mol CaF_2
   d. 1.35×10^{-4} mol Pb^{2+}
7. Calculate the number of moles represented by each of the following quantities.
   a. 3.11×10^{24} molecules of NO_2
   b. 8.06×10^{21} atoms of Pt
8. How many iodine atoms are in 2.45 mol of BaI_2?
9. Calculate the molar masses of the following substances.
   a. PCl_3
   b. $BaCO_3$
   c. $Fe_2(SO_4)_3$
   d. $Pb(CH_3COO)_2$
10. 3.50 moles of a certain compound contains 1.05×10^{25} carbon atoms. How many carbon atoms are in the formula of this compound?

Further Reading / Supplemental Links

- Kieffer, W.F., *Mole Concept in Chemistry*, Van Nostrand Reinhold Company, 1973.
- National Mole Day Foundation, INC., (http://www.moleday.org/)
- Molar Mass Calculator, (http://www.webpc.org/mmcalc.php)

Points to Consider

The molar mass of a compound can be used to convert between the mass of a substance (in grams) and its amount (in moles).

- What would the conversion factors for these calculations look like?
- Is there a way to convert between moles and the volume of a substance?

1.2 Mass, Volume, and the Mole

Lesson Objectives

- Use molar mass to make conversions between mass and moles of a substance.
- Explain Avogadro's hypothesis and how it relates to the volume of a gas at standard temperature and pressure.
- Convert between moles and volume of a gas at STP.
- Calculate the density of gases at STP.
- Use the mole road map to make two-step conversions between mass, number of particles, and gas volume.

Lesson Vocabulary

- Avogadro's hypothesis
- molar volume
- standard temperature and pressure (STP)

Check Your Understanding

Recalling Prior Knowledge

- What is a mole? How is the mole related to the number of particles in a sample of matter?
- What is the molar mass of a substance and how is it calculated?

Chemistry is about the study of chemical reactions. Quantitatively, chemicals react with one another in very specific ratios based upon the number of reacting particles. A more in depth study of chemical reactions requires the ability to make conversions between mass, volume, and moles.

Mole-Mass Relationship

You now know that the molar mass of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used to make calculations in the laboratory. Suppose that you need 3.00 moles of calcium chloride ($CaCl_2$) for a certain experiment. Since calcium chloride is a solid (**Figure** 1.5), it would be convenient to measure out this amount by using a balance, but, to do so, you would need to know what mass of $CaCl_2$ is equivalent to 3.00 moles. The molar mass of $CaCl_2$ is 110.98 g/mol, so a conversion factor can be constructed based on the fact that 1 mol $CaCl_2$ = 110.98 g $CaCl_2$. Dimensional analysis will then allow you to calculate the mass of $CaCl_2$ that you should use.

$$3.00 \text{ mol CaCl}_2 \times \frac{110.98 \text{ g CaCl}_2}{1 \text{ mol CaCl}_2} = 333 \text{ g CaCl}_2$$

When you measure out 333 g of $CaCl_2$, the resulting sample will contain 3.00 moles of $CaCl_2$.

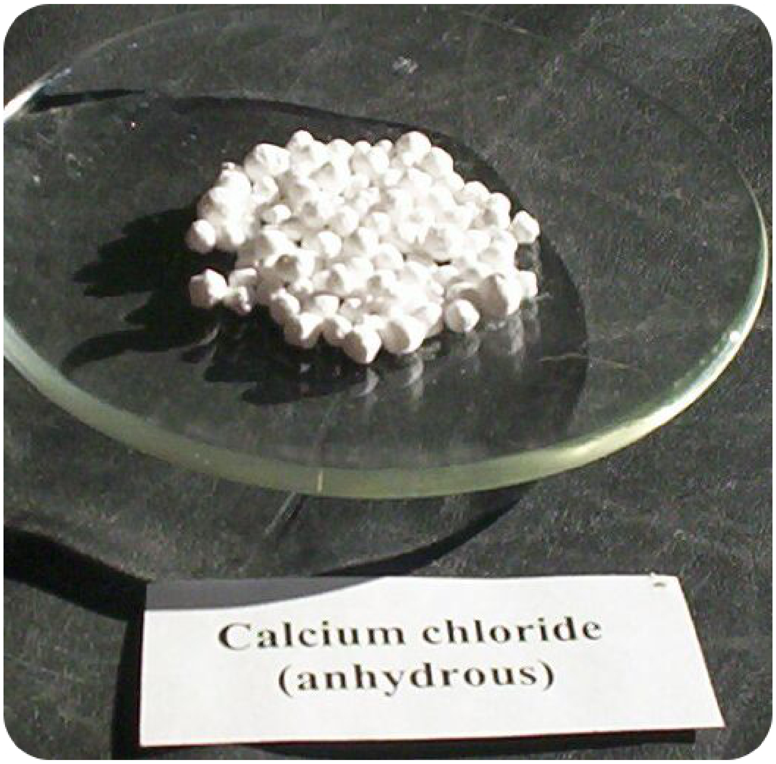


FIGURE 1.5
Calcium chloride is used as a drying agent and as a road deicer.

Sample Problem 10.4: Converting Moles to Mass

Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.

Step 1: List the known quantities and plan the problem.

Known

- 0.560 mol Cr
- molar mass of Cr = 52.00 g/mol

Unknown

- 0.560 mol Cr = ? g

The molar mass of chromium will allow us to convert from moles of Cr to grams.

Step 2: Calculate.

$$0.560 \text{ mol Cr} \times \frac{52.00 \text{ g Cr}}{1 \text{ mol Cr}} = 29.1 \text{ g Cr}$$

Step 3: Think about your result.

Since the desired amount was slightly more than one half of a mole, the mass should be slightly more than one half of the molar mass. The answer has three significant figures because the given value (0.560 mol) also has three significant figures.

Practice Problem

1. Find the masses of the following amounts.
   a. 2.15 mol of hydrogen sulfide, H_2S
   b. 3.95×10^{-3} mol of lead(II) iodide, PbI_2

A similar conversion factor based on molar mass can be used to convert the mass of a known substance to moles. In a laboratory situation, you might perform a reaction and produce a certain amount of a product. It will often be necessary to then determine the number of moles of the product that was formed, but this cannot be measured directly. However, you can use a balance to measure the mass of the product, and the number of moles can be easily calculated. The next Sample Problem illustrates this situation.

Sample Problem 10.5: Converting Mass to Moles

A certain reaction produces 2.81 g of copper(II) hydroxide, $Cu(OH)_2$. Determine the number of moles produced in the reaction.

Step 1: List the known quantities and plan the problem.

Known

- mass of $Cu(OH)_2$ produced = 2.81 g

Unknown

- amount of $Cu(OH)_2$ produced in moles

One conversion factor will allow us to convert from mass to moles.

Step 2: Calculate.

First, it is necessary to calculate the molar mass of $Cu(OH)_2$ from the molar masses of Cu, O, and H. The molar mass is 97.57 g/mol.

$$2.81 \text{ g Cu(OH)}_2 \times \frac{1 \text{ mol Cu(OH)}_2}{97.57 \text{ g Cu(OH)}_2} = 0.0288 \text{ mol Cu(OH)}_2$$

Step 3: Think about your result.

The molar mass is approximately 100 g/mol, so a quick estimate can be obtained by dividing the original value by 100 (moving the decimal point two places to the left). The relatively small mass of product formed results in a small number of moles.

Practice Problem

2. Calculate the number of moles represented by the following masses.
   1. 2.00×10^2 g of silver
   2. 37.1 g of silicon dioxide, SiO_2

Conversions Between Mass and Number of Particles

In the last lesson, "The Mole Concept," you learned how to convert back and forth between moles and the number of representative particles. Now you have seen how to convert back and forth between moles and the mass of a substance in grams. We can also combine these two types of problems. The figure below (**Figure** 1.6) illustrates that mass, number of particles, and moles are all interrelated. In order to convert between mass and number of particles, a conversion to moles is required first.

FIGURE 1.6
A conversion from number of particles to mass or from mass to number of particles requires two conversion factors.

Sample Problem 10.6: Converting Mass to Particles

How many molecules are present in a 20.0 g sample of chlorine gas, Cl_2?

Step 1: List the known quantities and plan the problem.

Known

- sample mass = 20.0 g Cl_2
- molar mass of Cl_2 = 70.90 g/mol

Unknown

- number of molecules of Cl_2

Use two conversion factors. The first converts grams of Cl_2 to moles. The second converts moles of Cl_2 to the number of molecules.

Step 2: Calculate.

$$20.0 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.90 \text{ g } Cl_2} \times \frac{6.02 \times 10^{23} \text{ molecules } Cl_2}{1 \text{ mol } Cl_2} = 1.70 \times 10^{23} \text{ molecules } Cl_2$$

The problem is done using two consecutive conversion factors. There is no need to explicitly calculate the moles of Cl_2.

Step 3: Think about your result.

Since the given mass is less than half of the molar mass of chlorine, the resulting number of molecules is less than half of Avogadro's number.

Practice Problems

3. How many formula units are contained in 270.2 g of zinc nitrate, $Zn(NO_3)_2$?
4. What is the mass of 5.84 $\times$ 10^{21} atoms of xenon?

Mole-Volume Relationship

Avogadro's Hypothesis and Molar Volume

In addition to number of particles and total mass, volume offers a third way to measure the amount of matter in a sample. With liquids and solids, the volume of a given number of particles can vary greatly depending on the density of the substance. This is because solid and liquid particles are packed close together with very little space in between. However, gases are largely composed of empty space between the actual gas particles (**Figure** 1.7).

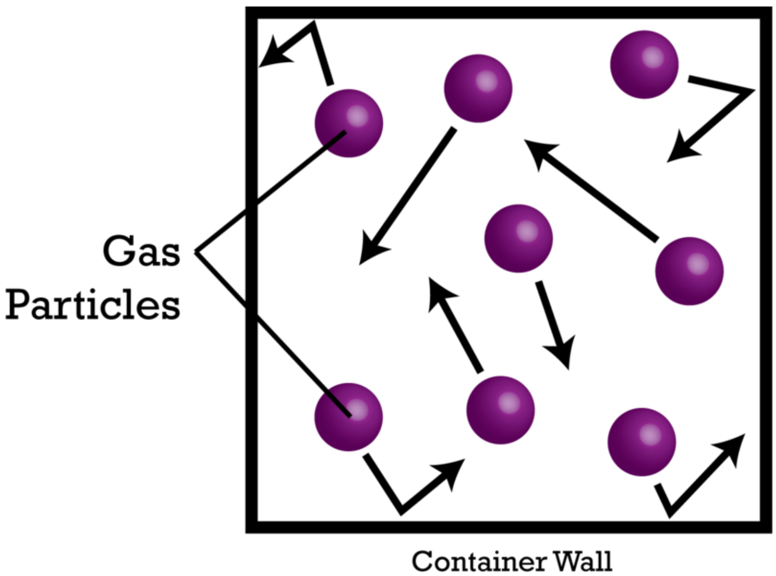


FIGURE 1.7
Gas particles are very small compared to the large amounts of empty space between them.

In 1811, Amedeo Avogadro suggested that the amount of gas in a given volume can be easily determined. **Avogadro's hypothesis** *states that equal volumes of all gases at the same temperature and pressure contain equal numbers of particles*. Since the total volume that a gas occupies is primarily composed of the empty space between the particles, the actual size of the particles themselves is nearly negligible. A given volume of a gas with small light particles, such as hydrogen (H_2), contains the same number of particles as the same volume of a heavy gas with larger particles, such as sulfur hexafluoride (SF_6).

Gases are compressible, meaning that when put under high pressure, the particles are forced closer to one another. This decreases the amount of empty space and reduces the volume of the gas. Gas volume is also affected by temperature. When a gas is heated, its molecules move faster and the gas expands. Because of the variation in gas volume due to pressure and temperature changes, gas volumes must be compared at the same temperature and pressure. **Standard temperature and pressure (STP)** *is defined as 0°C (273.15 K) and 1 atm of pressure*. The **molar volume** *of a gas is the volume of one mole of the gas at a given temperature and pressure*. At STP, one mole (6.02×10^{23} representative particles) of any gas occupies a volume of 22.4 L (**Figure** 1.8).

The figure below (**Figure** 1.9) illustrates how molar volume can be seen when comparing different gases. The given samples of helium (He), nitrogen (N_2), and methane (CH_4) are at STP. Each bulb contains 1 mole, or 6.02×10^{23} particles. However, because the gases have different molar masses (4.00 g/mol for He, 28.0 g/mol for N_2, and 16.0 g/mol for CH_4), the mass of each sample is different, even though each has the same number of particles.

You can watch a video experiment in which the molar volume of hydrogen gas at STP is determined at http://www.youtube.com/watch?v=6dmtLj2dLi0 .

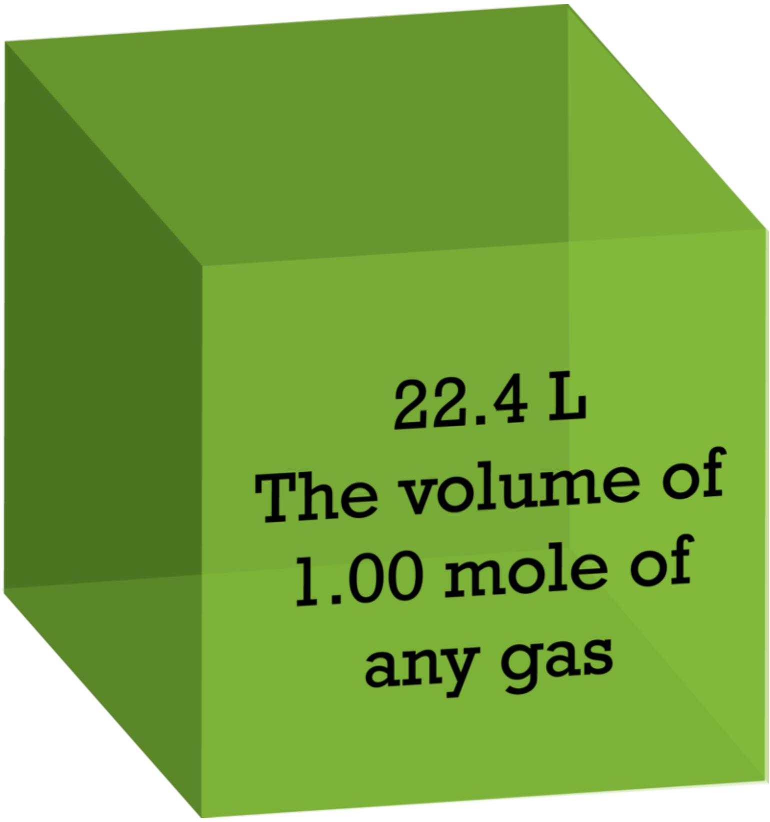


FIGURE 1.8
Any gas occupies 22.4 L at standard temperature and pressure (0 ℃ and 1 atm).

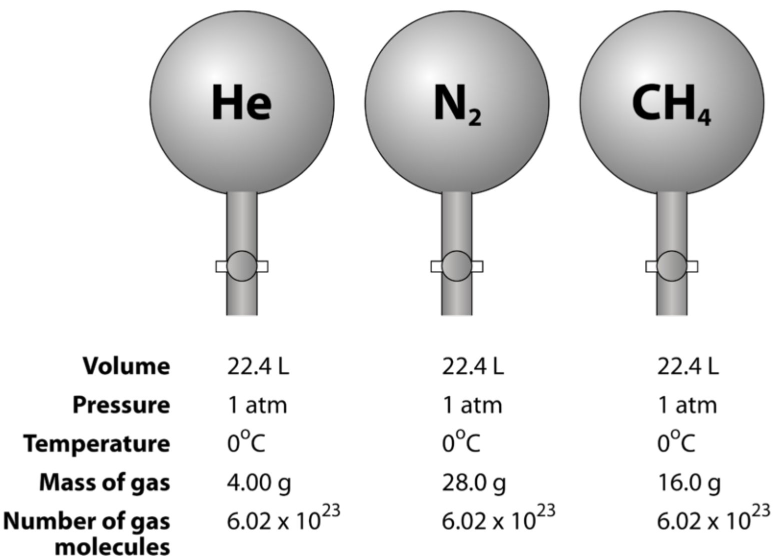
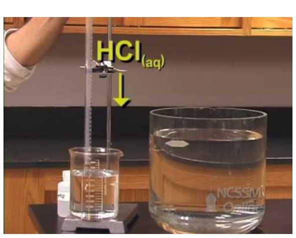


FIGURE 1.9
Avogadro's hypothesis states that equal volumes of any gas at the same temperature and pressure contain the same number of particles. At standard temperature and pressure, 1 mole of any gas occupies 22.4 L.



MEDIA
Click image to the left or use the URL below.
URL: http://www.ck12.org/flx/render/embeddedobject/60857

The lab document for this video can be found at http://www.dlt.ncssm.edu/core/Chapter7-Gas_Laws/Chapter7-Labs/Mg-HCl_Lab_web_01-02.doc .

Conversions Between Moles and Gas Volume

Molar volume at STP can be used to convert from moles to volume and from volume to moles for gaseous samples. The fact that 1 mole = 22.4 L is the basis for the conversion factor. This equality is only true at STP, so be sure that those are the specified conditions before using this as a conversion factor.

Sample Problem 10.7: Converting Gas Volume to Moles

Many metals react with acids to produce hydrogen gas. A certain reaction produces 86.5 L of hydrogen gas at STP. How many moles of H_2 were produced?

Step 1: List the known quantities and plan the problem.

Known

- volume of product = 86.5 L H_2
- 1 mol = 22.4 L

Unknown

- moles of H_2

Use the molar volume to convert from liters to moles.

Step 2: Calculate.

$$86.5 \text{ L } H_2 \times \frac{1 \text{ mol } H_2}{22.4 \text{ L } H_2} = 3.86 \text{ mol } H_2$$

Step 3: Think about your result.

The volume of gas produced is nearly four times larger than the molar volume. The fact that the gas is hydrogen plays no role in the calculation.

Practice Problems

5. How many moles of gas are present in 57.20 L of argon at a pressure of 1 atm and a temperature of 0°C?
6. At STP, what is the volume in milliliters of 0.0395 mol of fluorine gas, F_2?

Gas Density

As you know, density is defined as the mass per unit volume of a substance. Since gases all occupy the same volume on a per mole basis, the density of a particular gas at a given temperature and pressure is dependent only on its molar mass. A gas with a small molar mass will have a lower density than a gas with a large molar mass (**Figure** 1.10). Recall that gas densities are typically reported in g/L. Gas density can be calculated by combining molar mass and molar volume.

Sample Problem 10.8: Gas Density

FIGURE 1.10
Balloons filled with helium gas float in air because the density of helium is less than the density of air.

What is the density of nitrogen gas at STP?

Step 1: List the known quantities and plan the problem.

Known

- molar mass of N_2 = 28.02 g/mol
- 1 mol = 22.4 L

Unknown

- density of N_2 = ? g/L

Molar mass divided by molar volume yields the gas density at STP.

Step 2: Calculate.

$$\frac{28.02 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 1.25 \text{ g/L}$$

When these two ratios are multiplied in this way, the mol unit cancels, leaving g/L as the units for the answer.

Step 3: Think about your result.

The molar mass of nitrogen is slightly larger than molar volume, so its density is slightly greater than 1 g/L.

Alternatively, the molar mass of a gas can be determined if the density of the gas at STP is known.

Sample Problem 10.9: Molar Mass from Gas Density

What is the molar mass of a gas whose density is 0.761 g/L at STP?

Step 1: List the known quantities and plan the problem.

Known

- density = 0.761 g/L
- 1 mol = 22.4 L

Unknown

- molar mass = ? g/mol

Molar mass is equal to density (in g/L) multiplied by molar volume.

Step 2: Calculate.

$$\frac{0.761\ \text{g}}{1\ \text{L}} \times \frac{22.4\ \text{L}}{1\ \text{mol}} = 17.0\ \text{g/mol}$$

Step 3: Think about your result.

Because the density of the gas is less than 1 g/L, the molar mass is less than 22.4 g/mol.

Practice Problems

7. What is the density of sulfur dioxide, SO_2, at STP?
8. The density of an unknown noble gas is measured to be 1.78 g/L at STP. Calculate the molar mass and identify the noble gas.

Mole Road Map

Previously, we saw how the conversions between mass and number of particles required two steps, with moles as the intermediate. This concept can now be extended to also include gas volume at STP. The diagram shown below (**Figure** 1.11) is referred to as a mole road map.

The mole is the at the center of any calculation involving the amount of a substance. Sample Problem 10.10 is one of many different problems that can be solved using the mole road map.

Sample Problem 10.10: Mole Road Map

What is the volume of 79.3 g of neon gas at STP?

Step 1: List the known quantities and plan the problem.

Known

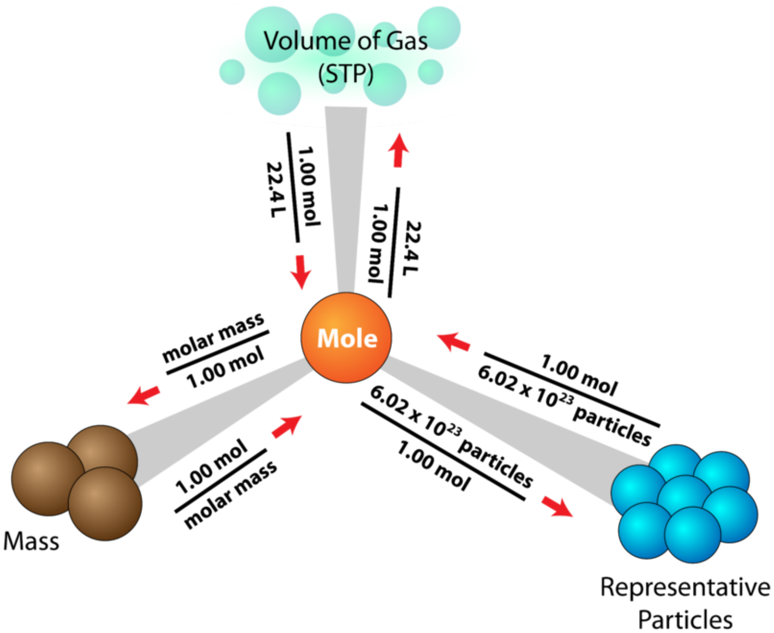


FIGURE 1.11
The mole road map shows the conversion factors needed to interconvert between mass, number of particles, and volume of a gas at STP.

- molar mass of Ne = 20.18 g/mol
- 1 mol = 22.4 L

<u>Unknown</u>

- sample volume = ? L

This problem can be solved by converting grams $\rightarrow$ moles $\rightarrow$ gas volume.

Step 2: Calculate.

$$79.3 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} \times \frac{22.4 \text{ L Ne}}{1 \text{ mol Ne}} = 88.0 \text{ L Ne}$$

Step 3: Think about your result.

The given mass of neon is equal to about 4 moles, resulting in a volume that is about 4 times larger than molar volume.

Lesson Summary

- The molar mass of a substance is used to convert grams to moles and moles to grams. Mass can be converted to the number of representative particles and vice-versa by using a two-step process.
- Avogadro's hypothesis states that equal volumes of all gases at the same temperature and pressure contain the same number of particles. The volume of 1 mole of any gas is called its molar volume and is equal to 22.4 L at standard temperature and pressure. Molar volume allows conversions to be made between moles and volume of gases at STP.
- Gas density can be calculated from the molar mass and molar volume.
- The mole road map outlines various possible conversions between mass, number of representative particles, and gas volume.

Lesson Review Questions

Reviewing Concepts

1. What do you need to know in order to calculate the number of moles of a substance from its mass?
2. Atoms of xenon gas are much larger than atoms of helium gas. Explain why the volume of 1 mole of xenon is the same as the volume of 1 mole of helium.
3. Why does the temperature and pressure need to be specified when working with the molar volume of a gas?
4. How would you expect the molar volume of a gas to change (increase or decrease) with the following changes in conditions?
    a. The temperature is increased.
    b. The external pressure exerted on the gas is increased.

Problems

5. Given the following two quantities: 0.50 mol of CH_4 and 1.0 mol of HCl,
    a. Which has more atoms?
    b. Which has more molecules?
    c. Which has the greater mass?
    d. Which has the greater volume at the same temperature and pressure (both are gases)?
6. How many moles of each substance are present in the following samples?
    a. 61.3 g of HBr
    b. 19.8 g of BeF_2
    c. 265 g of $AgNO_3$
    d. 0.412 kg of O_2
    e. 513 L of CO_2 gas at STP
    f. 1300. mL of He gas at STP
7. What is the mass in grams of each of the following?
    a. 3.20 mol of magnesium
    b. 6.55×10^{-3} mol of $(NH_4)_3PO_4$
    c. 12.20 mol of $SnSO_4$
    d. 4.05×10^{23} atoms of mercury
    e. 6.13×10^{24} molecules of I_2
    f. 15.4 L of N_2O gas at STP
    g. 247 L of C_3H_8 gas at STP
8. Determine the volume of the following gas quantities at STP.
    a. 2.78 mol of He
    b. 0.315 mol of CH_4
    c. 212 g of N_2
    d. 8.91 g of OF_2
    e. 3.36×10^{21} molecules of NH_3
    f. 7.81×10^{23} atoms of Kr
9. Make the following conversions.
    a. 612 g of Zn to atoms
    b. 18.77 L of CO gas (at STP) to molecules
    c. 2.10 g of $Ba(NO_3)_2$ to formula units

d. 25.0 mL of Ne gas (at STP) to atoms

10. What is the density in g/L of each of the following gases at STP?
    a. Cl_2
    b. He
11. A certain gas has a density of 2.054 g/L at STP. Calculate its molar mass. If the gas is known to consist of only nitrogen and oxygen, what is its formula?
12. Determine the number of C, H, and O atoms in 50.0 g of sucrose, $C_{12}H_{22}O_{11}$.
13. The density of aluminum metal is 2.70 g/cm^3. How many aluminum atoms are present in a cube of aluminum that measures 1.50 cm on each side?

Further Reading / Supplemental Links

- Gram/Mole/Volume Conversions, (http://www.sciencegeek.net/Chemistry/taters/Unit4GramMoleVolume.htm)

Points to Consider

The chemical formula of an ionic compound is an empirical formula, the simplest ratio between cations and anions in the crystal. The chemical formula of a molecular compound shows the number of each atom present in the compound.

- How is the composition of a compound related to its formula?
- How can mole calculations be used to analyze chemical formulas?

1.3 Chemical Formulas

Lesson Objectives

- Calculate the percent composition of a compound either from mass data or from the chemical formula. Use percent composition to calculate the mass of an element in a certain sample of a compound.
- Calculate the percentage of a hydrate's mass that is due to water.
- Determine the empirical formula of a compound from percent composition data.
- Determine the molecular formula of a compound from the empirical formula and the molar mass.

Lesson Vocabulary

- hydrate
- percent composition

Check Your Understanding

Recalling Prior Knowledge

- How is the mass of an element or compound converted to moles?
- What is an empirical formula? What is a molecular formula, and how does it relate to an empirical formula?

In previous chapters, you have learned about chemical nomenclature –naming compounds and writing correct chemical formulas. In this lesson, you will learn how the subscripts in a chemical formula represent the molar ratio between the elements in a compound.

Percent Composition

Packaged foods that you eat typically have nutritional information provided on the label. The label of a popular brand of peanut butter (**Figure** 1.12) reveals that one serving size is considered to be 32 g. The label also gives the masses of various types of compounds that are present in each serving. One serving contains 7 g of protein, 15 g of fat, and 3 g of sugar. This information can be used to determine the composition of the peanut butter on a percent by mass basis. For example, to calculate the percent of protein in the peanut butter, we could perform the following calculation:

$$\frac{7 \text{ g protein}}{32 \text{ g}} \times 100\% = 22\% \text{ protein}$$

FIGURE 1.12

Foods like peanut butter provide nutritional information on the label in the form of masses of different types of compounds present per serving.

In a similar way, chemists often need to know what elements are present in a compound and in what percentages. The **percent composition** *is the percent by mass of each element in a compound.* It is calculated in a way that is similar to what we just saw for the peanut butter.

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

Percent Composition from Mass Data

Sample Problem 10.11 shows how the percent composition of a compound can be calculated based on mass data.

Sample Problem 10.11: Percent Composition from Mass

A certain newly synthesized compound is known to contain the elements zinc and oxygen. When a 20.00 g sample of the compound is decomposed, 16.07 g of pure zinc remains. Determine the percent composition of the compound.

Step 1: List the known quantities and plan the problem.

Known

- total mass of sample = 20.00 g
- mass of Zn in the sample = 16.07 g

Unknown

- percent Zn = ? %
- percent O = ? %

First, subtract the mass of the zinc from the total mass to find the mass of oxygen in the sample. Then, divide each element's mass by the total mass to find the percent by mass.

Step 2: Calculate.

Mass of oxygen = 20.00 g –16.07 g = 3.93 g O

$$\% \text{ Zn} = \frac{16.07 \text{ g Zn}}{20.00 \text{ g}} \times 100\% = 80.35\% \text{ Zn}$$
$$\% \text{ O} = \frac{3.93 \text{ g O}}{20.00 \text{ g}} \times 100\% = 19.65\% \text{ O}$$

Step 3: Think about your result.

The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound is mostly zinc.

Practice Problems

1. A sample of a given compound contains 13.18 g of carbon and 3.32 g of hydrogen. What is the percent composition of this compound?
2. 5.00 g of aluminum is reacted with 7.00 g of fluorine to form a compound. When the compound is isolated, its mass is found to be 10.31 g, with 1.69 g of aluminum left unreacted. Determine the percent composition of the compound.

Percent Composition from a Chemical Formula

The percent composition of a compound can also be determined from its chemical formula. The subscripts in the formula are first used to calculate the mass of each element found in one mole of the compound. That value is then divided by the molar mass of the compound and multiplied by 100%.

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

The percent composition of a given compound is always the same as long as the compound is pure.

Sample Problem 10.12: Percent Composition from a Chemical Formula

Dichlorine heptoxide (Cl_2O_7) is a highly reactive compound used in some synthesis reactions. Calculate the percent composition of dichlorine heptoxide.

Step 1: List the known quantities and plan the problem.

Known

- mass of Cl in 1 mol Cl_2O_7 = 70.90 g
- mass of O in 1 mol Cl_2O_7 = 112.00 g
- molar mass of Cl_2O_7 = 182.90 g/mol

Unknown

- percent Cl = ? %
- percent O = ? %

Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the compound by the molar mass of the compound and multiplying by 100%.

Step 2: Calculate.

$$\% \text{ Cl} = \frac{70.90 \text{ g Cl}}{182.90 \text{ g}} \times 100\% = 38.76\% \text{ Cl}$$
$$\% \text{ O} = \frac{112.00 \text{ g O}}{182.90 \text{ g}} \times 100\% = 61.24\% \text{ O}$$

Step 3: Think about your result.

The percentages add up to 100%.

Practice Problem

3. Calculate the percent composition of the following compounds:
    1. magnesium fluoride, MgF_2
    2. silver nitrate, $AgNO_3$

Percent composition can also be used to determine the mass of a certain element that is contained in a sample whose total mass is known. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is 38.76% Cl and 61.24% O. Suppose that you needed to know the masses of chlorine and oxygen present in a 12.50 g sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element.

$$12.50 \text{ g } Cl_2O_7 \times \frac{38.76 \text{ g Cl}}{100 \text{ g } Cl_2O_7} = 4.845 \text{ g Cl}$$
$$12.50 \text{ g } Cl_2O_7 \times \frac{61.24 \text{ g O}}{100 \text{ g } Cl_2O_7} = 7.655 \text{ g O}$$

The sum of the two masses is 12.50 g, the mass of the total sample.

Percent of Water in a Hydrate

Many ionic compounds naturally contain water as part of the crystal lattice structure. A **hydrate** *is a compound that has one or more water molecules bound to each formula unit*. Ionic compounds that contain a transition metal are often highly colored. Interestingly, it is common for the hydrated form of a compound to be of a different color than the anhydrous form, which has no water in its structure. A hydrate can usually be converted to its anhydrous form by heating. The figure below (**Figure** 1.13) shows that the anhydrous compound cobalt(II) chloride is blue, while its hydrate is a distinctive magenta color.

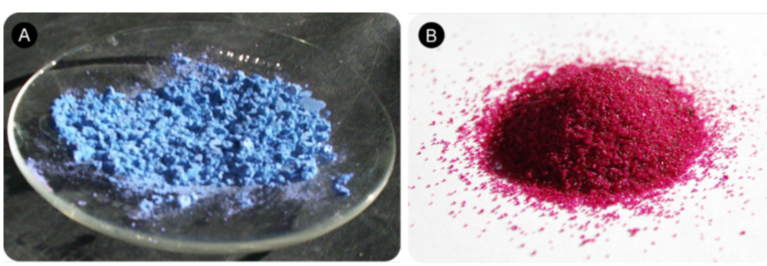


FIGURE 1.13

On the left is anhydrous cobalt(II) chloride ($CoCl_2$). On the right is the hydrated form of the compound, cobalt(II) chloride hexahydrate ($CoCl_2$•6 H_2O).

The hydrated form of cobalt(II) chloride contains six water molecules in each formula unit. The name of the compound is cobalt(II) chloride hexahydrate, and its formula is $CoCl_2$•6 H_2O. The formula for water is set apart at the end of the formula with a dot, preceded by a coefficient that represents the number of water molecules per formula unit.

It is sometimes useful to know what percent of a hydrate's mass is water. Sample problem 10.13 demonstrates the procedure for finding this value.

Sample Problem 10.13: Percent of Water in a Hydrate

What percent of cobalt(II) chloride hexahydrate ($CoCl_2$•6 H_2O) is water?

Step 1: List the known quantities and plan the problem.

The mass of water in one mole of the hydrate is the coefficient (6) multiplied by the molar mass of H_2O. The molar mass of the hydrate is the molar mass of $CoCl_2$ plus the mass of the associated water.

Known

- mass of H_2O in 1 mole of hydrate = 108.12 g
- molar mass of hydrate = 237.95 g/mol

Unknown

- percent H_2O = ? %

Calculate the percent by mass of water by dividing the mass of H_2O in 1 mole of the hydrate by the molar mass of the hydrate and multiplying by 100%.

Step 2: Calculate.

$$\% \ H_2O = \frac{108.12 \text{ g } H_2O}{237.95 \text{ g}} \times 100\% = 45.44\% \ H_2O$$

Step 3: Think about your result.

Nearly half of the mass of the hydrate is from water molecules within the crystal.

Practice Problem

4. Gypsum is a soft mineral used in plaster and is composed of calcium sulfate dihydrate. Calculate the percent by mass of water in calcium sulfate dihydrate, $CaSO_4$•2 H_2O.

Empirical Formulas

Recall that an empirical formula is one that shows the lowest whole-number ratio of the elements in a compound. Because the structure of ionic compounds is an extended three-dimensional network of positive and negative ions, only empirical formulas are used to describe ionic compounds. However, we can also consider the empirical formula of a molecular compound. Ethene is a small hydrocarbon compound with the formula C_2H_4 (**Figure** 1.14). While C_2H_4 is its molecular formula and represents its true molecular structure, it has an empirical formula of CH_2. The simplest ratio of carbon to hydrogen in ethene is 1:2. In each molecule of ethene, there is 1 carbon atom for every 2 atoms of hydrogen. Similarly, we can also say that in one mole of ethene, there is 1 mole of carbon for every 2 moles of hydrogen. The subscripts in a formula represent the molar ratio of the elements in that compound.

In a procedure called elemental analysis, an unknown compound can be analyzed in the laboratory to determine the percentages of each element contained within it. These values can be used to find the molar ratios of the elements, which gives us the empirical formula. The steps to be taken are outlined below.

1. Assume a 100 g sample of the compound so that the given percentages can be directly converted into grams.
2. Use each element's molar mass to convert the grams of each element to moles.

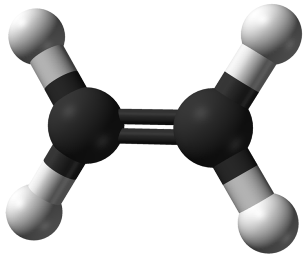


FIGURE 1.14
Ball-and-stick model of ethene, C_2H_4.

3. In order to find a whole-number ratio, divide the moles of each element by the smallest value obtained in step 2.
4. If all the values at this point are whole numbers (or very close), each number is equal to the subscript of the corresponding element in the empirical formula.
5. In some cases, one or more of the values calculated in step 3 will not be whole numbers. Multiply each of them by the smallest number that will convert all values into whole numbers. Note that all values must be multiplied by the same number so that the relative ratios are not changed. These values can then be used to write the empirical formula.

Sample Problem 10.14: Determining the Empirical Formula of a Compound

A compound of iron and oxygen is analyzed and found to contain 69.94% iron and 30.06% oxygen. Find the empirical formula of the compound.

Step 1: List the known quantities and plan the problem.

Known

- % of Fe = 69.94%
- % of O = 30.06%

Unknown

- Empirical formula = $Fe_?O_?$

Follow the steps outlined in the text.

Step 2: Calculate.

1. Assume a 100 g sample.

$\rightarrow$ 69.94 g Fe

$\rightarrow$ 30.06 g O

2. Convert to moles.

$$69.94 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 1.252 \text{ mol Fe}$$
$$30.06 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.879 \text{ mol O}$$

3. Divide both values by the smallest of the results.

$$\frac{1.252 \text{ mol Fe}}{1.252} = 1 \text{ mol Fe}$$
$$\frac{1.879 \text{ mol O}}{1.252} = 1.501 \text{ mol O}$$

4. Since the moles of O is still not a whole number, both numbers can be multiplied by 2. The results are now close enough to be rounded to the nearest whole number.

$$1 \text{ mol Fe} \times 2 = 2 \text{ mol Fe}$$
$$1.501 \text{ mol O} \times 2 = 3 \text{ mol O}$$

The empirical formula of the compound is Fe_2O_3.

Step 3: Think about your result.

The subscripts are whole numbers and represent the molar ratio of the elements in the compound. The unknown compound is iron(III) oxide.

Practice Problem

5. Calculate the empirical formula of each compound from the percentages listed.
    1. 63.65% N, 36.35% O
    2. 81.68% C, 18.32% H

Molecular Formulas

Molecular formulas tell us how many atoms of each element are present in one molecule of a molecular compound. In many cases, the molecular formula is the same as the empirical formula. For example, the molecular formula of methane is CH_4, and, because 1:4 is the smallest whole-number ratio that can be written for this compound, that is also its empirical formula. Sometimes, however, the molecular formula is a simple whole-number multiple of the empirical formula. Acetic acid is an organic acid that gives vinegar its distinctive taste and smell. Its molecular formula is $C_2H_4O_2$. Glucose is a simple sugar that cells use as their primary source of energy. Its molecular formula is $C_6H_{12}O_6$. The structures of both molecules are shown below (**Figure** 1.15). They are very different compounds, yet both have the same empirical formula, CH_2O.

Empirical formulas can be determined from the percent composition of a compound. In order to determine its molecular formula, it is necessary to also know the molar mass of the compound. Chemists use an instrument called a mass spectrometer to determine the molar mass of compounds. In order to go from the empirical formula to the molecular formula, follow these steps:

1. Calculate the empirical formula mass (EFM), which is simply the molar mass represented by the empirical formula.
2. Divide the molar mass of the compound by the empirical formula mass. The result should be a whole number or very close to a whole number.
3. Multiply all of the subscripts in the empirical formula by the whole number found in step 2. The result is the molecular formula.

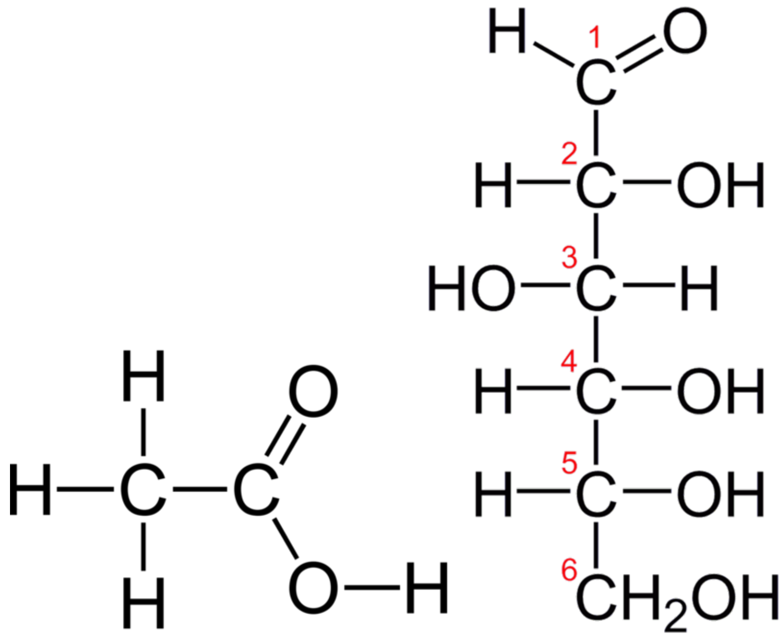


FIGURE 1.15
Acetic acid (*left*) has a molecular formula of $C_2H_4O_2$, while glucose (*right*) has a molecular formula of $C_6H_{12}O_6$. Both have the empirical formula CH_2O.

Sample Problem 10.15: Determining the Molecular Formula of a Compound

The empirical formula of a compound that contains boron and hydrogen is BH_3. Its molar mass is 27.7 g/mol. Determine the molecular formula of the compound.

Step 1: List the known quantities and plan the problem.

Known

- empirical formula = BH_3
- molar mass = 27.7 g/mol

Unknown

- molecular formula = ?

Step 2: Calculate.

1. The empirical formula mass (EFM) = 13.84 g/mol
2. $\frac{\text{molar mass}}{\text{EFM}} = \frac{27.7}{13.84} = 2$
3. $BH_3 \times 2 = B_2H_6$

The molecular formula of the compound is B_2H_6.

Step 3: Think about your result.

The molar mass of the molecular formula matches the molar mass of the compound.

Practice Problems

6. A compound with the empirical formula CH has a molar mass of 78 g/mol. Determine its molecular formula.
7. A compound is found to consist of 43.64% phosphorus and 56.36% oxygen. The molar mass of the compound is 284 g/mol. Find the molecular formula of the compound.

You can watch a video lecture about molecular and empirical formulas at http://www.khanacademy.org/science/chemistry/chemical-reactions-stoichiometry/v/molecular-and-empirical-formulas .

You can watch a video lecture about determining molecular and empirical formulas from percent composition at http://www.khanacademy.org/science/physics/thermodynamics/v/molecular-and-empirical-forumlas-from-percent-composition .

Lesson Summary

- The percent composition of a compound is the percent by mass of each of the elements in the compound. It can be calculated from mass data or from the chemical formula.
- A hydrate is an ionic compound that contains one or more water molecules in the crystal lattice for each formula unit. The percentage of a hydrate's mass that is composed of water can be calculated from its formula.
- Percent composition data can be used to determine a compound's empirical formula, which is the molar ratio between the elements in the compound.
- The empirical formula and the molar mass of a substance can be used to determine its molecular formula, which is the number of each kind of atom in a single molecule of the compound.

Lesson Review Questions

Reviewing Concepts

1. How many moles of aluminum, sulfur, and oxygen atoms are there in one mole of $Al_2(SO_4)_3$?
2. How can a hydrate be converted into its anhydrous form?
3. Give the empirical formula for each of the following molecular compounds.
   a. C_8H_{18}
   b. H_2O_2
   c. N_2O_4
   d. C_5H_{12}
4. What is the relationship between a compound's empirical formula and its molecular formula?

Problems

5. A sample of a certain binary compound contains 6.93 g of silicon and 7.89 g of oxygen. What is the percent composition of the compound?
6. Calculate the percent composition for each of the following compounds.
   a. potassium bromide, KBr
   b. ammonium chloride, NH_4Cl

   c. acetone, C_3H_6O
   d. barium phosphate, $Ba_3(PO_4)_2$

7. Using the answers from problem 6, calculate the mass of the indicated element in each of the following samples.
   a. potassium in 4.23 g of KBr
   b. chlorine in 126 g of NH_4Cl
   c. carbon in 41.0 g of C_3H_6O
   d. phosphorus in 39.6 g of $Ba_3(PO_4)_2$
8. Which compound has the highest nitrogen content by mass?
   a. KNO_3
   b. NO_2
   c. NH_4Cl
   d. Li_3N
9. Find the percentage of the total mass that is due to water for each of the following hydrates.
   a. $ZnSO_4$•7 H_2O
   b. $Mn(NO_3)_2$•4 H_2O
10. Find the empirical formulas of compounds with the following percent compositions.
   a. 38.35% Cl and 61.65% F
   b. 59.35% Sr, 8.135% C, and 32.51% O
11. Find the molecular formula of each compound, given its empirical formula and molar mass.
   a. CH_2O, 120 g/mol
   b. C_2HCl, 181.5 g/mol
12. The molar mass of a compound is 92 g/mol. Analysis of a sample of the compound indicates that it contains 0.606 g N, 1.390 g O, and no other elements. Find its molecular formula.
13. 12.50 g of a hydrated form of copper(II) sulfate, $CuSO_4$•x H_2O (where x is unknown) is gently heated. When all the water has been driven off, the mass of the anhydrous copper(II) sulfate is found to be 7.99 g.
   a. What is the mass of the water that was lost as a result of the heating?
   b. Convert this mass of water to moles of water.
   c. Convert the mass of the anhydrous $CuSO_4$ to moles.
   d. Divide the answer to b by the answer to c. This is the value of x in the formula of the hydrate. Write the formula of hydrated copper(II) sulfate.

Further Reading / Supplemental Links

You can watch video lectures about formulas from mass composition at https://www.khanacademy.org/science/chemistry/chemical-reactions-stoichiometry/v/formula-mass-composition and another mass composition problem at http://www.khanacademy.org/science/chemistry/chemical-reactions-stoichiometry/v/another-mass-composition-problem .

Points to Consider

Chemical reactions are the essence of chemistry. We will describe chemical reactions both qualitatively and quantitatively.

- What are some ways to classify different types of chemical reactions?
- How are calculations with moles, grams, and volume involved in the analysis of chemical reactions?

1.4 References

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