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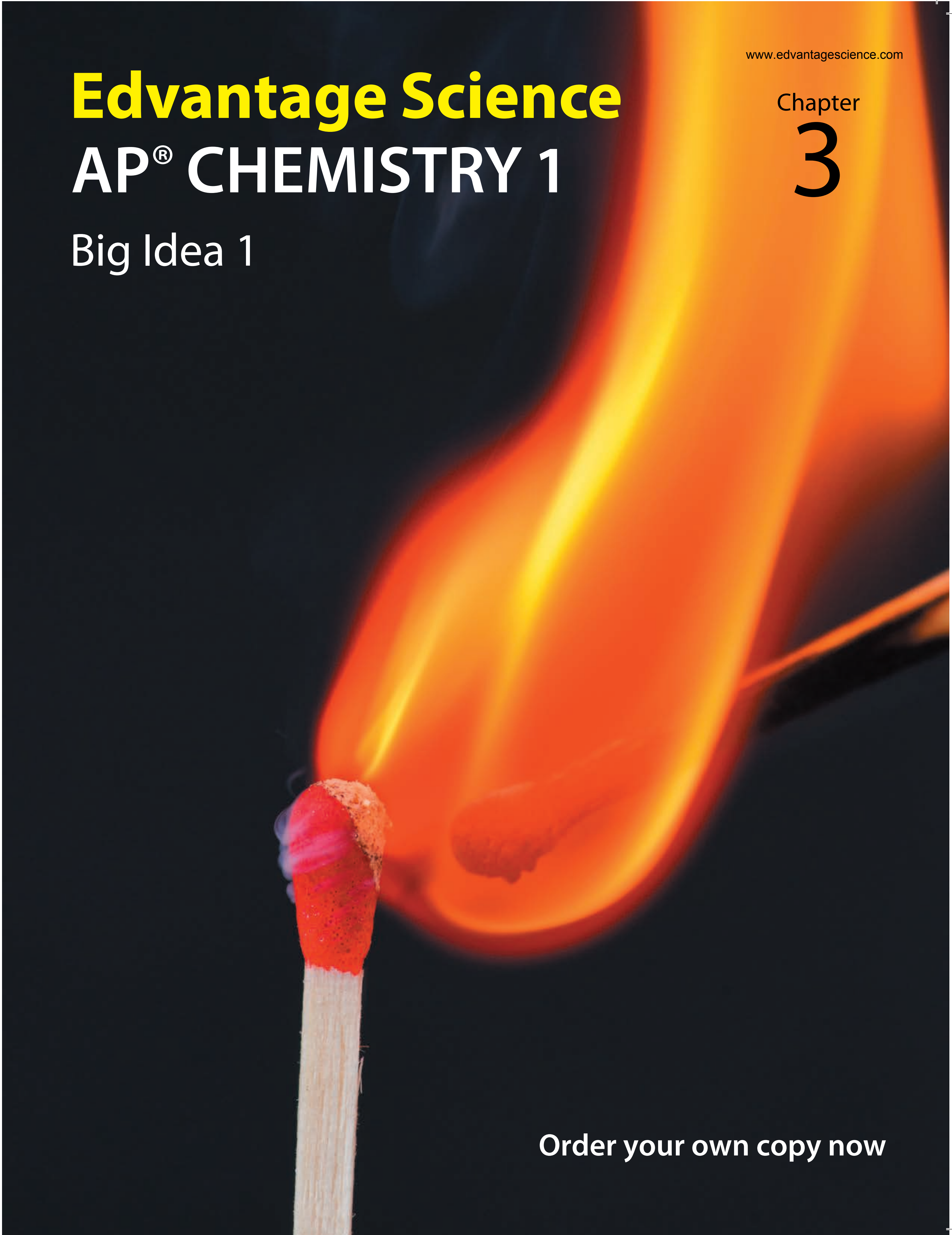
# AP® CHEMISTRY 1

## Big Idea 1

Chapter 3



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*AP Chemistry 1*



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# 3 The Mole — The Central Unit of Chemistry

This chapter focuses on the following AP Big Idea from the College Board:

**Big Idea 1:** The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identify in chemical reactions.

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By the end of this chapter, you should be able to do the following:

- Explain the significance and use of the mole
- Perform calculations involving the mole
- Determine relationships between molar quantities of gases at STP
- Perform calculations involving molecular and empirical formulas to identify a substance
- Describe concentration in terms of molarity
- Perform calculations involving molarity

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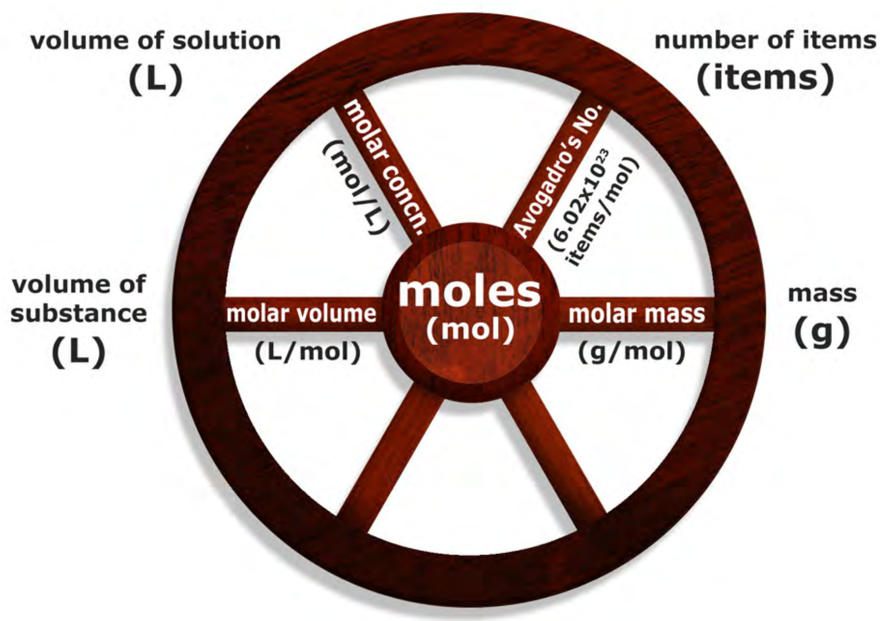
By the end of this chapter you should know the meaning of these **key terms**:

- empirical formula
- molarity
- molar mass
- molar solution
- molar volume
- mole
- molecular formula
- molecular mass
- percentage composition
- relative atomic mass
- standard solution
- stoichiometry
- STP



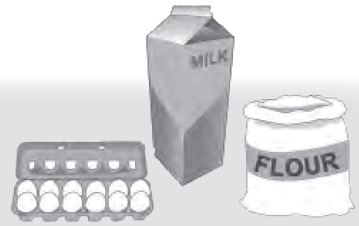

*The mole is at the centre of the chemical measurement.*

# 3.1 Relative Atomic Mass

## Warm Up

1. Complete the grocery list by filling in the missing units (Figure 3.1a).
   One __________ eggs, Two __________milk, Three _________flour
2. From your answer to #1, what are the three ways that we typically express amounts of materials?

   (a) number of items (b) ____________________ (c) __________________




*We express amounts of materials in different ways.*

### Relative Mass

Experimental chemistry is essentially figuring out things about matter that cannot be observed directly. The joy of experimental chemistry lies in figuring out how to figure it out. Consider the following ingenious method for determining the relative masses of objects without needing to know their actual masses. **Mass** is the amount of matter. When you say that one object has twice as much mass as another you are expressing the object's **relative mass**. You are comparing one object's mass to the other's.

Suppose you wanted to determine the relative mass of a staple and a grain of rice, each of which is too small to register a mass on your balance (Figure 3.1.1). Why not weigh 100 of each? If 100 identical staples weigh twice as much as 100 identical grains of rice then one staple will weigh twice as much as one grain of rice. The nifty aspect of this technique is that we don't even need to know how many objects we are weighing: we just need to know that we're weighing the same number of each. If some number of staples weighs twice as much as the same number of rice grains then any number of staples will weigh twice as much as that number of rice grains.

This technique for determining relative mass still works even if the items being weighed are not identical. If the items being weighed for comparison are not identical, then the ratio provided is that of their average masses rather than the ratio of the masses of the individual items since this would depend on which individual items. For example, if a variety of pens weighs 1.52 times as much as the same number of a variety of pencils then the average mass of these pens is 1.52 times the average mass of these pencils.


**Figure 3.1.1** *One hundred staples and one hundred grains of rice have the same mass ratio as one staple and one grain of rice, 2:1.*

The mass ratio of any equal number of items equals the average mass ratio of those individual items.

While you should never confuse the terms "weight" and "mass," the word "weigh" serves double duty. To **weigh** is to find the weight or compare the weights of. Since scales work by comparing weights, you are by definition "weighing" objects and materials with a scale. In fact, a **weighing scale** is a measuring instrument for determining the mass or weight of an object.

### Quick Check

1. What does "relative mass" mean? ____________________________________________
2. You have two bags of candy from a bulk food store: a bag of gumdrops and a bag of jujubes. You intend to determine the relative masses of a jujube and a gumdrop by weighing the contents of each bag. What condition is necessary for this to work? ________________________________________

## Law of Constant Composition

In this book, as in most chemistry textbooks, much of our current chemical knowledge will be presented in historical context. Instead of just telling you what we know (or think we know), we'll tell you how chemists came to this understanding. This is because chemistry is more than just an accumulated list of facts about matter: it is also the processes that lead us to such information. By learning and assessing these processes, as well as the facts, some of you will decide to continue this quest. In addition, people often acquire a better understanding of a concept by learning the concept in the same manner that it was originally developed.

To use the technique just described to determine the relative masses of different types of atoms, chemists needed to be able to weigh an equal number of different types of atoms. In the early 1800s, chemists discovered that all samples of a given compound have the same mass ratio of their constituent elements. For example, there are 8 g of oxygen for every 1 g of hydrogen in every sample of water. This is called the **law of constant composition**. In 1804, John Dalton, a scientist in England, argued that the law of constant composition not only supported the concept of atoms but also provided their relative masses. He reasoned that the mass ratios in which different elements combine are the mass ratios of their individual atoms or a simple multiple thereof. If one atom of magnesium weighs 1.5 times as much as one atom of oxygen then any number of magnesium atoms would weigh 1.5 times as much as the same number of oxygen atoms. Dalton argued that this was the reason all samples of a compound contained the same mass ratio of its elements.

### Sample Problem — Determining Relative Atomic Mass

A chemist carefully heats 0.350 g of magnesium powder in a crucible. The magnesium reacts with atmospheric oxygen to produce 0.580 g of magnesium oxide (MgO). What is the mass of a magnesium atom relative to the mass of an oxygen atom?

| What to Think about | How to Do It |
| --- | --- |
| 1. 0.350 g Mg must have combined with 0.230 g O to produce 0.580 g MgO. | 0.580 g MgO – 0.350 g Mg = 0.230 g O |
| 2. Since magnesium oxide has the formula MgO, 0.350 g of magnesium and 0.230 g of oxygen contain equal numbers of atoms. | $\frac{\text{mass of Mg atoms}}{\text{mass of O atoms}} = \frac{0.350 \text{ g}}{0.230 \text{ g}} = 1.52$ |
| 3. If some number of Mg atoms weighs 1.52 times as much as the same number of O atoms then any number of Mg atoms weighs 1.52 times as much as the same number of O atoms, even one of each. | A Mg atom weighs 1.52 times as much as an O atom. |

### Practice Problems — Determining Relative Atomic Mass

1. A dozen identical AA batteries have a mass of 276 g and a dozen identical watch batteries have a mass of only 26.4 g. The mass of an AA battery is _______ times the mass of a watch battery.

2. A sample of strontium oxide (SrO) is found to contain 2.683 g Sr and 0.490 g O. What is the mass of a strontium atom relative to that of an oxygen atom?

3. A 4.218 g sample of daltonium bromide (DBr) is decomposed and 0.337 g of D is recovered.
   (a) What is the atomic mass of daltonium given that the atomic mass of bromine is 79.9 u?
   ______________________________________________

   (b) This question uses the fictitious element, daltonium, so you can't just look up the element's atomic mass.
   What element does daltonium represent? ______________________________

## Relative Masses of Atoms

According to the sample problem, if all the atoms of an element are identical then the mass of a magnesium atom is 1.52 times the mass of an oxygen atom. If all the atoms of an element do not have the same mass then the average mass of a magnesium atom is 1.52 times the average mass of an oxygen atom. The issue of whether all the atoms of an element are identical wasn't resolved for another century but, as described, we need only insert the word, "average" if they are not.

The element hydrogen was discovered to have the least massive atoms so its atoms were originally assigned an atomic mass of 1 u (atomic mass unit) and the mass of all the other types of atoms were expressed relative to this. The discussion of atomic mass and atomic mass units will continue in chapter 5. Oxygen's atomic mass of 16 u means that the mass of an oxygen atom is 16 times the mass of a hydrogen atom (or that the average mass of an oxygen atom is 16 times the average mass of a hydrogen atom) (Figure 3.1.2). If the mass of a magnesium atom is 1.52 times the mass of an oxygen atom then the mass of a magnesium atom is $1.52 \times 16.0$ u = 24.3 u. The periodic table of the elements confirms that magnesium has a relative atomic mass of 24.3 u.

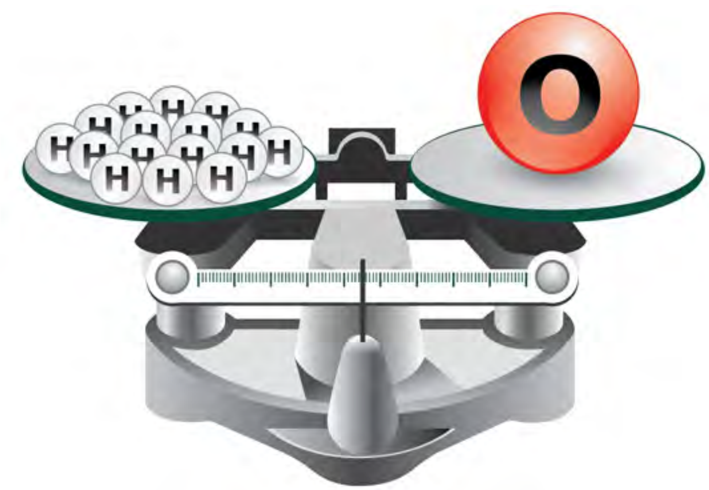



**Figure 3.1.2** *The mass of an oxygen atom is equal to the mass of 16 hydrogen atoms.*

Determining the relative masses of the basic units of matter was a remarkable feat. Dalton bridged the gap between the world we experience and the invisible world of atoms by deriving the relative masses of atoms from laboratory observations. But how did Dalton know that the formula of magnesium oxide was MgO? Recall Dalton's important qualification: "or a simple multiple thereof." If the formula of magnesium oxide is $MgO_2$ then the mass ratio of Mg to O in the compound would need to be doubled to determine their atomic mass ratio. This is necessary because we are weighing half as many magnesium atoms. Therefore, the same number of magnesium atoms would weigh twice as much. Similar adaptations would be required for other possible formulas.

### Sample Problem — Determining Relative Atomic Mass (Non 1:1 Formulas)

Barium chloride has a mass ratio of 1.934 g Ba:1.000 g Cl. Chlorine has an atomic mass of 35.5 u. What is the atomic mass of barium if the formula of barium chloride is $BaCl_2$?

| What to Think about | How to Do It |
| --- | --- |
| If the formula is $BaCl_2$ then we need to double the mass of barium so that we can compare the masses of equal numbers of atoms. | $2(1.934) \times 35.5$ u = 137.3 u |

### Practice Problems — Determining Relative Atomic Mass (Non 1:1 Formulas)

Aluminum iodide has a mass ratio of 1.000 g Al: 14.100 g I. Given that the atomic mass of iodine is 126.9 u, what is the atomic mass of aluminum if the formula of aluminum iodide is:

1. $AlI_3$? ____________________  2. $Al_2I_3$? ____________________

## Cannizzaro's Paper

Dalton assumed that atoms combined in the simplest manner possible. He believed that if a pair of elements (A + B) formed only one compound, the formula for the compound would be AB. If they formed a second compound, its formula would be either $A_2B$ or $AB_2$. Dalton was well aware that he had no evidence for his "rules of simplicity." He conceded that some of his formulas and resulting atomic mass determinations might be incorrect. As you may recall from Science 9 and Science 10, the formulas of ionic compounds are simple ratios, but not quite as simple as Dalton supposed.

On September 3, 1860, many of Europe's leading chemists met in Karlsruhe, Germany. At this meeting, the Italian chemist Stanislao Cannizzaro presented a remarkable paper in which he solved the mystery of atomic masses. For example, Dalton hadn't understood how two particles of hydrogen gas could react with one particle of oxygen gas to produce two particles of water vapour. He thought that couldn't happen because it would require splitting the oxygen particle, which he thought was an atom. Cannizzaro showed that Dalton's atomic model was still valid if the hydrogen and oxygen gas particles were made up of pairs of atoms. Hydrogen and oxygen molecules are called **diatomic molecules** because they are formed of two atoms of the same element ("di" means 2).

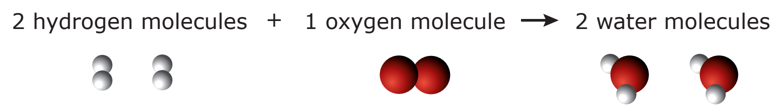
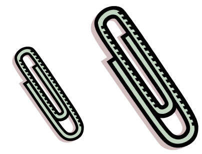



**Figure 3.1.3** *Diatomic molecules of hydrogen and oxygen combine to form water molecules.*

Cannizzaro's paper went on to describe and explain three other techniques for determining atomic mass: one for metals, one for liquid or gaseous non-metals, and one for solid non-metals.

Dalton is called the father of the atomic theory because he explained how the law of constant composition provided support for the concept of atoms. However, additional methods were required to determine the relative atomic masses. These atomic masses were, in turn, used to determine the correct formulas of compounds. Dmitri Mendeleev, who published his first periodic table of the elements in 1869, was at Karlsruhe. The correct atomic masses were a prerequisite to Mendeleev's famous table.

# 3.1 Activity: The Relative Mass of Paper Clips



### Question
What is the mass of a large paper clip relative to that of a small paper clip?
(We'll answer this question without weighing only one paper clip of either type.)

### Background
If some number of large paper clips weighs twice as much as the same number of small paper clips then any number of large paper clips will weigh twice as much as the same number of small paper clips, including one of each. Remember we don't need to know how many paper clips we are weighing; we just need to know that we're weighing the same number of each. The mass ratio of any equal number of identical items equals the mass ratio of the individual items.

### Procedure
1. Weigh a pile of small paper clips. Record this mass in the table provided below.
2. Attach a large paper clip to each small paper clip and measure the total mass of these coupled clips. Record this mass in the table provided below.
3. Calculate the total mass of the attached large paper clips and record this mass in the table below.

### Results and Discussion

| Objects | Mass (g) |
|---|---|
| Small paper clips | |
| Coupled paper clips | |
| Large paper clips | |

1. $\dfrac{\text{mass of some number of large paper clips}}{\text{mass of the same number of small paper clips}} = \dfrac{\_\_\_\_\_\_\_\_\_\_\ \text{g}}{\_\_\_\_\_\_\_\_\_\_\ \text{g}} = \_\_\_\_\_\_\_\_\_\_$

   The mass of a large paper clip is ________ times the mass of a small paper clip.

2. If we assign a small paper clip a mass of 1.00 smu (stationary mass unit), what is the mass of a large paper clip?

3. Let's check this result by weighing one small paper clip and one large paper clip.

   $\dfrac{\text{mass of one large paper clip}}{\text{mass of one small paper clip}} = \dfrac{\_\_\_\_\_\_\_\_\_\_\ \text{g}}{\_\_\_\_\_\_\_\_\_\_\ \text{g}} = \_\_\_\_\_\_\_\_\_\_$

4. Why might the ratios calculated in steps 1 and 3 be slightly different?

## 3.1 Review Questions

1. A certain number of identical glass marbles has a mass of 825 g. The same number of identical steel marbles has a mass of 2245 g.
   (a) Assigning a glass marble a mass of 1.00 mmu (marble mass unit), calculate the mass of a steel marble.

   (b) Why don't you need to know the number of marbles that were weighed?

2. 1.965 g of sodium is placed in a flask containing chlorine gas. 5.000 g of NaCl is produced in the resulting reaction.
   (a) A sodium atom's mass is _______ times a chlorine atom's mass.

   (b) Chlorine has an atomic mass of 35.5 u. What is the atomic mass of sodium?

3. A 10.000 g sample of zubenium fluoride (ZuF) is decomposed and 8.503 g of Zu is recovered.
   (a) What is the atomic mass of zubenium?

   (b) This question uses the fictitious element zubenium so you can't just look up the element's atomic mass. What element does zubenium represent?

4. Zinc sulphide has a mass ratio of 2.037 g Zn: 1.000 g S. Given that the atomic mass of sulphur is 32.1 u, what is the atomic mass of zinc if the formula of zinc sulphide is:
   (a) ZnS?

   (b) $ZnS_2$?

   (c) $Zn_3S_2$?

5. A compound of copper and oxygen contains 13.073 g Cu and 1.647 g O. Oxygen has an atomic mass of 16.0 u.
   (a) What is the atomic mass of copper if the formula of the above compound is CuO?

   (b) What is the atomic mass of copper if the formula of the above compound is $Cu_2O$?

   (c) What is the atomic mass of copper if the formula of the above compound is $CuO_2$?

6. In 1819, Dulong and Petit noted a relationship between the presumed atomic mass of most metals and their specific heats. The specific heat of a metal divided into 25.0 provides the approximate atomic mass of the metal. The specific heat of a substance is the amount of heat required to raise 1 g of the substance by 1°C. The specific heat of copper is 0.3864 J/g°C.
   (a) Calculate the approximate atomic mass of copper using Dulong and Petit's method.

   (b) Knowing the approximate atomic mass of the metal allowed chemists to determine which of the more accurate atomic masses derived by composition analysis was correct. Which of the atomic masses and corresponding formulas calculated in question 5 is correct for the compound that was analyzed?

7. Determine the percent error of Dulong and Petit's method of approximating a metal's atomic mass for aluminum (0.903 J/g°C), magnesium (1.05 J/g°C) and silver (0.23772 J/g°C).

8. In 1811 Amedeo Avogadro proposed that equal volumes of any gas at the same temperature and pressure contain *the same number of particles*. Cannizzaro realized this allows scientists to weigh equal numbers of atoms of different gaseous elements and determine their relative atomic masses. Complete the following data table showing the mass of equal volumes of two different gases at the same temperature and pressure.

| Element | Mass of Gas (g) | Relative Atomic Mass (u) |
|---|---|---|
| H | 0.210 | 1.0 |
| | 7.455 | |

9. Potassium has an atomic mass of 39.1 u. What does this mean?

10. Look up the following elements in the periodic table and report each element's atomic mass.

(a) P ______________

(b) Ca ______________

(c) U ______________

11. Eight identical forks have a mass of 213.1 g. Eight identical knives have a mass of 628.2 g.

(a) What is the mass of a knife relative to that of a fork?

(b) Why did you not need to divide the supplied masses by 8 to answer 10(a)?

(c) What could you conclude from this data if the utensils of each type were not identical?

12. A mint is advertising a special set of silver coins containing a 10 g coin, a 20 g coin and a 30 g coin. One of these coins is accidentally being made 1 g lighter than its advertised mass. You have two sets of these coins and have been challenged to identify the undersized coin by weighing only one pile of coins. The single pile may include any combination of the coins that you wish. What combination of the coins would you weigh? How can you use that mass to identify the undersized coin?

# 3.2 Introducing the Mole — The Central Unit of Chemistry

## Warm Up

1. It takes 15 gulps to drink a bottle of water. What information would allow you to calculate how many *slurps* it would take to drink an identical bottle of water?
   ______________________
2. If 4 slurps = 1 gulp, how many slurps would it take to consume a 15-gulp drink?
   ____________________________________________________________
3. If 5 slurps equal 1 gulp, how many gulps would it take to consume a 20-slurp drink? ___________

## The Mole Concept

What mass of oxygen has the same number of atoms as 1 g of hydrogen? An oxygen atom (16 u) weighs 16 times as much as a hydrogen atom (1 u). Therefore, it would require 16 g of oxygen to have the same number of atoms as 1 g of hydrogen. Chemists extended this reasoning to all the elements. For example, 55.8 g Fe, 35.5 g Cl, 23.0 g Na, and 12.0 g C all contain the same number of atoms since these masses are in the same ratios as their individual atomic masses. How many atoms are there in the atomic mass of any element expressed in grams? Originally chemists didn't know and even now they only have a very rough estimate but they nevertheless gave a name to that number. They called this number a "mole."

> A **mole** is a quantity equal to the number of atoms in the atomic mass of any element expressed in grams (e.g., the number of atoms in 1.0 g H, 16.0 g O, 63.5 g Cu).

The definition of the mole is under continuous review. It is "fine-tuned" periodically in response to new information about atomic structure and to changes in the definition of the atomic mass unit on which the definition of the mole is based.


**Figure 3.2.1** *$6.02214179 \times 10^{23}$ carbon atoms*

The number of things in a mole is also referred to as **Avogadro's number** in honour of the Italian scientist whose insight regarding gases led to a technique for determining the relative atomic masses of non-metals. Just as the word "dozen" refers to a number of something, so does the word "mole." The chief difference is that we know that a dozen is 12 of something but we only have a rough estimate of how many things are in a mole of something. There have been many independent derivations of the number of items in a mole. Chemists currently estimate that a mole is $6.02214179 \times 10^{23}$ give or take a few million billion (Figure 3.2.1). The actual number isn't important unless you're working at the atomic level because whatever the number is, it's the same for a mole of anything.

Just as a dozen is 12 of anything, a mole is approximately $6.02 \times 10^{23}$ of anything. While a dozen is a fairly small number, a mole is an absurdly large number. A mole of peas would cover the entire Earth's surface with a layer over 200 m deep. Atoms are so small however that you can hold a mole of atoms in the palm of your hand. Just as a dozen is a convenient unit of quantity for a baker to group buns and doughnuts, a mole is a convenient unit of quantity for a chemist to group atoms and molecules. The number of items is one way to express the amount of a material so chemists often refer to a mole of a substance rather than a mole of the substance's particles (e.g., a mole of copper instead of a mole of copper atoms).

The mole was introduced in the early 1900s by Wilhelm Ostwald. Ironically, Ostwald developed the mole concept as an alternative to the atomic theory, which he did not accept. Today, the mole is used throughout modern chemistry as the central unit through which all other quantities of materials are related, but it was not common before the mid-1950s, just two generations of chemists ago. Before that, chemists related quantities of chemicals through their atomic masses without reference to the mole.

## Quick Check

1. (a) How is a mole like a dozen?

   ______________________________________________

   (b) How is a mole different than a dozen?

   ______________________________________________

2. What does a mole of chlorine atoms weigh? ______________

3. What mass of sulphur has the same number of atoms that are in 1.0 g H? ______________

## Conversions Using Avogadro's Number

Many chemistry problems can be solved by examining the units of the values provided and considering the units that the answer must have. This technique of manipulating units is called "dimensional analysis." While it's important to understand the mathematical operation you're performing, it's reassuring to see that the units work out (i.e., that the answer has the appropriate units).

Most chemical conversions involve the mole. The key to conversion is the conversion factor. Chemists know or know where to find the conversion factors they need. At this point in the course, we'll use only two conversion factors: Avogadro's number (the number of items per mole) and molar mass (the number of grams per mole). The items that a chemist would normally be concerned about are chemical species such as atoms, molecules, ions, and formula units. The abbreviation for the unit mole is mol (Figure 3.2.2).

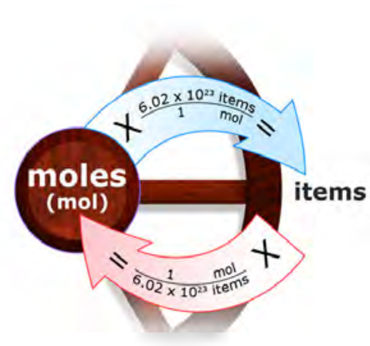



**Figure 3.2.2** *Converting the moles and the number of items of a substance*

| Name | Equivalence Statement | Conversion Factors | |
|---|---|---|---|
| Avogadro's number | 1 mol = $6.02 \times 10^{23}$ items | $\frac{6.02 \times 10^{23} \text{ items}}{1 \text{ mol}}$ | $\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ items}}$ |

### Sample Problem—Converting Moles to Number of Items

How many oxygen atoms are in 3.2 mol of oxygen atoms?

#### What to Think about

1. Convert: mol O ⟶ atoms O
2. Setup:
   $3.2 \text{ mol O} \times \frac{? \text{ atoms O}}{1 \text{ mol O}}$
3. Conversion factor:
   $6.02 \times 10^{23}$ atoms O per 1 mol O
4. Count the number of significant figures of each value in the operation and then round the answer to the least of these.

#### How to Do It

$$3.2 \cancel{\text{mol O}} \times \frac{6.02 \times 10^{23} \text{ atoms O}}{1 \cancel{\text{mol O}}} = 1.9 \times 10^{24} \text{ atoms O}$$

Note: There is no uncertainty in the 1 mol O. The uncertainty of the conversion factor is expressed in the $6.02 \times 10^{23}$ atoms O.

## Practice Problems—Converting Moles to Number of Items

1. Chromium ions are responsible for the beautiful colours of rubies and emeralds.
   How many chromium ions ($Cr^{3+}$) are in 3.5 mol of chromium ions?

2. 30.0 mol $H_2O$ = ____________ molecules $H_2O$

3. How many atoms of sodium are in 0.023 mol Na?

## Sample Problem—Converting Number of Items to Moles

$7.3 \times 10^{24}$ carbon monoxide molecules (CO) represent how many moles of carbon monoxide?

| What to Think about | How to Do It |
|---|---|
| 1. Convert: molecules CO $\rightarrow$ mol CO | |
| 2. Setup: $(7.3 \times 10^{24} \text{ molecules CO}) \times \frac{1 \text{ mol CO}}{? \text{ molecules CO}}$ | $(7.3 \times 10^{24}\ \cancel{\text{molecules CO}}) \times \frac{1 \text{ mol CO}}{6.02 \times 10^{23}\ \cancel{\text{molecules CO}}}$ |
| 3. Conversion factor: 1 mol CO per $6.02 \times 10^{23}$ molecules CO | $= 12 \text{ mol CO}$ |

## Practice Problems — Converting Number of Items to Moles

1. Incandescent lights are filled with argon to prevent the glowing filament from burning up.
   How many moles of argon do $1.81 \times 10^{22}$ atoms of argon represent?

2. $2.25 \times 10^{24}$ molecules $CO_2$ = ____________ mol $CO_2$?

3. A 1-L intravenous bag of saline solution contains $9.27 \times 10^{22}$ formula units of NaCl.
   How many moles of NaCl is this?

## Molar Mass

The mass of one mole of an element's atoms is called that element's **molar mass** (Figure 3.2.3). It follows from simply restating the definition of a mole that the molar mass of an element is its atomic mass expressed in grams. For example, "one mole is the number of atoms in 16 g of oxygen" can be restated as "one mole of oxygen atoms weighs 16 g." The atomic masses of the elements can be found in the periodic table. The atomic mass of oxygen is 16 u and thus the molar mass of oxygen is 16 g. This is better expressed as a conversion factor for calculation purposes: 16 g per mole of oxygen or 16 g/mol O.


**Figure 3.2.3** *The mass of 1 mol of a chemical depends on the atoms that make it up.*

The **molecular mass** or **formula mass** of a compound is the sum of its constituent atomic masses (e.g., $H_2O$: 2(1 u) + 16 u = 18 u). One mole of water molecules consists of 1 mol of oxygen atoms (16 g) and 2 mol of hydrogen atoms (2 g) and therefore weighs 18 g. Similarly, 1 mol of NaCl formula units consists of 1 mol of sodium atoms (23 g) and 1 mol of chlorine atoms (35.5 g) for a total mass of 58.5 g (Figure 3.2.4).

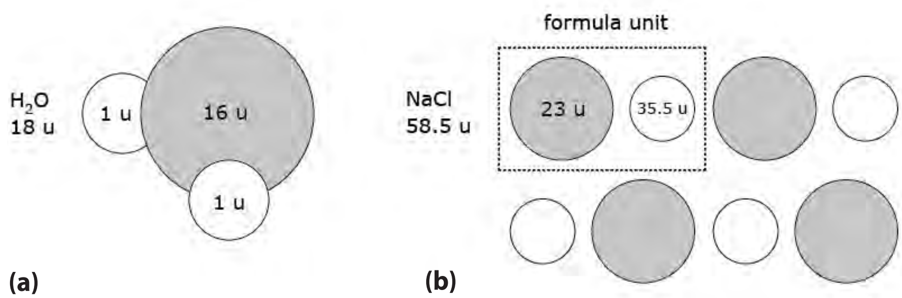



**Figure 3.2.4** *(a) The molecular mass of water is the sum of the masses of the oxygen and hydrogen atoms. (b) The formula mass of NaCl is the sum of the masses of sodium and chlorine atoms.*

Just as the molar mass of an element is simply its atomic mass expressed in grams, the **molar mass** of a compound is simply its molecular or formula mass expressed in grams.

> The molar mass of a substance is its atomic, molecular, or formula mass expressed in grams.

### Sample Problem — Determining a Compound's Formula Mass and/or Molar Mass

What are the formula mass and molar mass of $Al_2(SO_4)_3$?

| What to Think About | How to Do It |
|---|---|
| 1. 1 $Al_2(SO_4)_3$ consists of 2 Al's, 3 S's, and 12 O's. | Formula Mass = 2(27.0 u) + 3(32.1 u) + 12(16.0 u) = 342.3 u |
| 2. 1 mol $Al_2(SO_4)_3$ consists of 2 mol Al, 3 mol S and 12 mol O. | Molar Mass = 2(27.0 g) + 3(32.1 g) + 12(16.0 g) = 342.3 g<br>Expressed as a conversion factor, it is 342.3 g/mol. |

### Practice Problems — Determining a Compound's Formula Mass and/or Molar Mass

1. What is the molecular mass of nitrogen dioxide?

2. What is the molar mass of $Na_2Cr_2O_7$?

3. What is the molar mass of iron(III) sulphide?

## Conversions Using Molar Mass

| Name | Equivalence Statement | Conversion Factors | |
|---|---|---|---|
| Molar mass | 1 mol = ? g | $\frac{?\ \text{g}}{1\ \text{mol}}$ | $\frac{1\ \text{mol}}{?\ \text{g}}$ |
| Example: $H_2O$ | 1 mol = 18.0 g | $\frac{18.0\ \text{g}}{1\ \text{mol}}$ | $\frac{1\ \text{mol}}{18.0\ \text{g}}$ |

### Sample Problem — Converting Moles to Mass

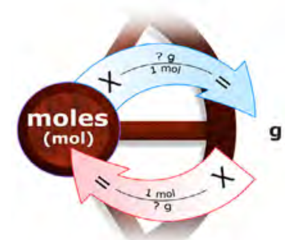
What is the mass of 3.2 mol of oxygen atoms?

**What to Think about**

1. Convert: mol O → g O
2. Setup:
   $3.2\ \text{mol O} \times \frac{?\ \text{g O}}{1\ \text{mol O}}$
3. Conversion factor:
   16.0 g O per 1 mol O

**How to Do It**

$$3.2\ \cancel{\text{mol O}} \times \frac{16.0\ \text{g O}}{1\ \cancel{\text{mol O}}} = 51\ \text{g O}$$




### Practice Problems — Converting Moles to Mass

1. What does 2.65 mol of table salt (NaCl) weigh?

2. 0.87 mol $NH_3$ = ________ g $NH_3$?

3. Very large quantities of chemicals are produced in the chemical industry. Worldwide production of sulphuric acid ($H_2SO_4$) is estimated at two trillion ($2.0 \times 10^{12}$) moles annually. How many tonnes of $H_2SO_4$ is this? (1 tonne = 1000 kg)

## Sample Problem — Converting Mass to Moles

How many moles of water are in 1.8 g of water?

| What to Think about | How to Do It |
|---|---|
| 1. Convert: g $H_2O \rightarrow$ mol $H_2O$<br>2. Setup:<br>$1.8\ g\ H_2O \times \frac{1\ mol\ H_2O}{?\ g\ H_2O}$<br>3. Conversion factor.<br>1 mol $H_2O$ per 18 g $H_2O$ | $1.8\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.0\ \cancel{g\ H_2O}} = 0.10\ mol\ H_2O$ |

## Practice Problems — Converting Mass to Moles

1. Gold is the most malleable metal. It can be hammered into sheets that are only several hundred atoms thick. In 2010, Vancouver's Science World covered an entire billboard with just two troy ounces (62.2 g) of gold to dramatize this fact. How many moles of gold is this?

2. 3.88 g $CO_2$ = __________ mol $CO_2$

3. Smelling salts are used to revive an unconscious athlete. A capsule of smelling salts contains 500.0 mg $(NH_4)_2CO_3$. How many moles of $(NH_4)_2CO_3$ is this?

# 3.2 Activity: A Mole of Pennies

## Question

How long and massive would a stack of 1 mol of pennies be?

## Background

The mole is a convenient and useful term for counting very large quantities of things. You know that 1 mol of pennies is approximately $6.02 \times 10^{23}$ pennies but can you picture just how many that really is? Suppose you stacked 1 mol of pennies. How tall do you think that stack would be? How much would it weigh? Just for fun, try guessing by completing the tables below in pencil before you do the necessary calculations.

## Procedure

1. Make a stack of 10 pennies.
2. Measure and record the stack's height in centimetres. _______________
3. Measure and record the stack's mass in grams. __________________

## Results and Discussion

1. Calculate the height in kilometres of a stack of 1 mol of pennies.

2.

| Would the stack reach… | Distance (km) | ✔ or ✗ |
|---|---|---|
| our Moon? | $3.9 \times 10^{5}$ | |
| Pluto? | $5.9 \times 10^{9}$ | |
| Proxima Centauri (the nearest star)? | $4.1 \times 10^{13}$ | |
| Andromeda (the nearest galaxy)? | $1.9 \times 10^{19}$ | |

3. Calculate the mass in kilograms of 1 mol of pennies.

4.

| Would the stack weigh as much as… | Mass (kg) | ✔ or ✗ |
|---|---|---|
| the U.S.S. *Ronald Reagan* (the world's heaviest aircraft carrier)? | $2.1 \times 10^{7}$ | |
| the total of all living things on Earth? | $2 \times 10^{15}$ | |
| our Moon? | $7.4 \times 10^{22}$ | |
| Earth? | $6.0 \times 10^{24}$ | |

## 3.2 Review Questions

1. (a) What is the definition of a mole?

   (b) What is our best estimate of the number of things in a mole?

   (c) What do chemists call this number?

2. (a) What mass of carbon would have the same number of atoms as 1.0 g H?

   (b) What mass of carbon would have the same number of atoms as 3.0 g H?

   (c) What mass of sulphur would have the same number of atoms as 32.0 g O?

3. (a) What does a mole of iron weigh?

   (b) Chemists call this value the ________________________ of iron.

4. (a) What is the molecular mass of propane, $C_3H_8$?

   (b) What is the formula mass of calcium hydroxide, $Ca(OH)_2$?

   (c) What is the molar mass of carbon tetrachloride, $CCl_4$?

5. 3.2 mol C = ____________________ atoms C

6. How many molecules are in 0.0085 mol of $C_2H_6$?

7. $1.4 \times 10^{18}$ Ag atoms represent how many moles of atoms?

8. 2.99 g Na = ___________ mol Na

9. What is the mass of 5.2 mol of fluorine?

10. Airline regulations prohibit lithium metal batteries that contain over 2.0 g of lithium on passenger aircraft. How many moles of lithium are in 2.0 g Li?

11. What is the mass of 0.32 mol of sodium nitrite?

12. A can of cola contains 58 mg of caffeine, $C_8H_{10}N_4O_2$. How many moles of caffeine are in a can of cola?

13. Carbon dioxide, produced by respiration in plants and animals, causes the slightly acidic nature of normal rain. How many molecules of $CO_2$ are in 0.725 mol $CO_2$?

14. The male luna moth can detect specialized chemicals known as pheromones in order to locate a mate. A moth can detect $1.70 \times 10^9$ molecules of the pheromone. How many moles of the pheromone is this?

15. Cycling enthusiasts often prefer bicycles made with titanium frames. Titanium is resistant to corrosion and fatigue, has a significantly lower density than steel, and seems to have a natural shock absorbing ability. Suppose a high-quality titanium frame contains 1300 g of titanium. How many moles of titanium does this frame contain?

16. Bluestone is an attractive mineral with the chemical name copper(II) sulphate pentahydrate. What is the mass of a 1.75-mol sample of bluestone?

17. An environmental assessment predicts that a coal plant would emit $8.18 \times 10^6$ mol of ammonia into the atmosphere annually. How many tonnes of ammonia is this?

18. Ammonium phosphate is a fertilizer containing nitrogen and phosphorus for healthy plant growth. How many moles of ammonium phosphate are in a bag containing 2.640 kg of it?

19. The movie *Erin Brockovich* dramatizes the efforts of the title character (played by Julia Roberts) to prove that the Pacific Gas and Electric Co. contaminated the water supply of Hinkley, California, with hexavalent chromium. Tin(II) dichromate is a hexavalent chromium compound. What is the mass of 5.925 mol of tin(II) dichromate?

# 3.3 The Wheel Model of Mole Conversions

## Warm Up

1. Which contains more atoms, 30 g Cl or 15 g C? ____________________
2. Which weighs more, 1 mol Zn or 3 mol N? ____________________
3. Which contains more molecules, 34 g $CH_4$ or 58 g $O_2$? ____________________

### Two-Step Mole Conversions

The mole serves as a link between the invisible world of atoms and observable quantities of chemicals. The mole is to a chemist what the dollar is to an accountant. Just as the dollar is the central unit of commerce and allows us to keep track of money, the mole is the central unit of chemistry and allows us to keep track of atoms and molecules.

A quantity of a substance can only be related to another quantity of the same substance through the mole.

How would you solve the following problem?
How many atoms are in 5.0 g of copper?

You might split the problem into two parts, each of which you learned how to solve in section 3.2:

1. How many moles of copper are in 5.0 g Cu?
2. How many copper atoms is this?

This is how chemists solve this type of problem. Think of the mole as the hub of a wheel with the spokes leading out to all the other units. In our wheel model, the spokes represent the conversion factors (Figure 3.3.1).

For now, our conversions are limited to those between moles and items and between moles and grams. Mass and the number of items must be related to each other through the mole: grams to moles to items or items to moles to grams. The beauty of the wheel model is that as you learn more chemical quantities they can simply be added to the rim of the wheel. In order to relate or "connect" a new chemical quantity to all of the others you only need to connect it to the mole. In other words, if you wanted to convert grams into sneebugs, you would convert grams into moles and then moles into sneebugs.

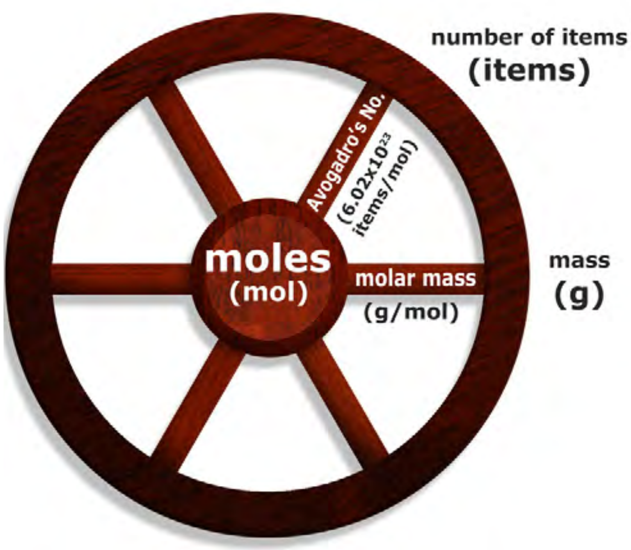



**Figure 3.3.1** *The wheel is a useful model for representing conversion factors.*

## Sample Problem — Two-Step Conversion: Mass to Number of Items (Atoms)

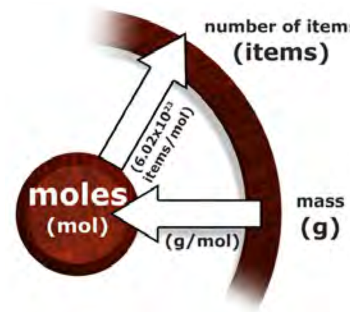
Chemists count by weighing. How many atoms are in 5.0 g of copper?

### What to Think about

1. Plan your route.
   Convert: g Cu → mol Cu → atoms Cu
   Chemists usually perform these two calculations in one continuous sequence.
2. Setup: $5.0 \text{ g Cu} \times \dfrac{1 \text{ mol Cu}}{? \text{ g Cu}} \times \dfrac{? \text{ atoms Cu}}{1 \text{ mol Cu}}$

   The numerators show the route; in this case, grams to moles to atoms.
   Each numerator's unit is cancelled by the next denominator's unit until you arrive at your answer.
3. Conversion factors:
   1 mol Cu per 63.5 g Cu
   $6.02 \times 10^{23}$ atoms Cu per 1 mol Cu

### How to Do It




$$5.0 \cancel{\text{ g Cu}} \times \frac{1 \cancel{\text{ mol Cu}}}{63.5 \cancel{\text{ g Cu}}} \times \frac{6.02 \times 10^{23} \text{ atoms Cu}}{1 \cancel{\text{ mol Cu}}}$$

$$= 4.7 \times 10^{22} \text{ atoms Cu}$$

As Avogadro's number is an estimate, so is the above answer. Nevertheless, being able to estimate the number of atoms in any sample of a substance is remarkable. This is perhaps even more evident for the reverse conversion, which allows us to estimate the mass of a single atom in grams.

## Sample Problem — Two-Step Conversion: Number of Items (Atoms) to Mass

What is the mass of an oxygen atom in grams?

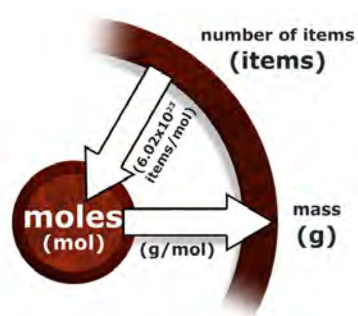
### What to Think about

1. Convert: atom O → mol O → g O
2. Setup: $1 \text{ atom O} \times \dfrac{1 \text{ mol O}}{? \text{ atoms O}} \times \dfrac{? \text{ g O}}{1 \text{ mol O}}$
3. Conversion factors:
   1 mol O per $6.02 \times 10^{23}$ atoms O
   16.0 g per 1 mol O
4. All that we've done here is divide the mass of a mole of oxygen atoms by the number of atoms in a mole to determine the mass of each atom.
   Always check to make sure that your answer is reasonable. $2.66 \times 10^{-23}$ g is a reasonable answer because an atom would have a very small mass.

### How to Do It

$$1 \cancel{\text{ atom O}} \times \frac{1 \cancel{\text{ mol O}}}{6.02 \times 10^{23} \cancel{\text{ atoms O}}} \times \frac{16.0 \text{ g O}}{1 \cancel{\text{ mol O}}}$$

$$= 2.66 \times 10^{-23} \text{ g O}$$



## Practice Problems — Two-Step Conversions

1. Fill in the missing entries to determine the mass in grams of a billion billion ($1 \times 10^{18}$) sulphur dioxide molecules.

$$(1 \times 10^{18} \text{ molecules } SO_2) \times \frac{1 \text{ mol } SO_2}{\_\_\_\_\_\_\_\_ \text{ molecules } SO_2} \times \frac{\_\_\_\_\_\_ \text{ g } SO_2}{1 \text{ mol } SO_2} = \_\_\_\_\_\_\_\_ \text{ g } SO_2$$

2. How many atoms are in 2.1 g Br?

3. What is the mass in grams of one atom of Ag?

### Two Wheel Conversions (Composition Stoichiometry)

Chemists often relate a quantity of one chemical substance to a quantity of another. **Stoichiometry** is the branch of chemistry that deals with the quantitative relationships between elements in a compound (composition stoichiometry) and between the reactants and products in a chemical reaction (reaction stoichiometry). ("Stoichiometry" is from the Greek words *stoicheion* meaning "element" and *metron* meaning "measure.") Our presentation here is limited to composition stoichiometry. To accommodate such conversions, we simply add another wheel to our model. Each wheel in our model represents a different substance or species and, of course, the only functional way to connect two wheels is with an axle. An axle runs between the hubs of wheels and in our model, connects moles of one substance or species to moles of another (Figure 3.3.2).

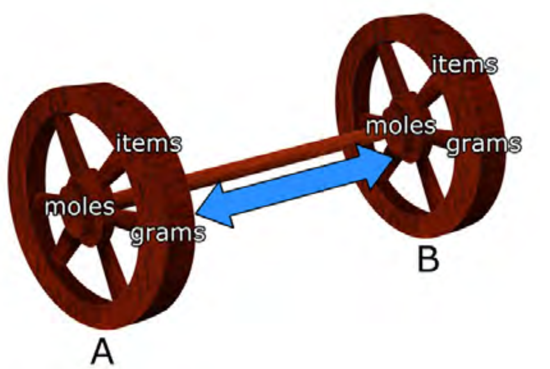



**Figure 3.3.2** *The two-wheel-and-axle model for converting between species*

> A quantity of a substance or species can only be related to a quantity of another substance or species through the mole.

We already know how to move about (convert units) using single wheels. The only new step added here is represented by the axle, which helps you to convert moles of one substance into moles of another. The conversion factor is found in the compound's formula. There are two oxygen atoms in a $CO_2$ molecule, so there are two dozen oxygen atoms in a dozen $CO_2$ molecules and there are… *wait for it*… 2 mol of oxygen atoms in 1 mol of $CO_2$ molecules.

| Name | Equivalence Statement | Conversion Factors | |
|---|---|---|---|
| Chemical Formula | 1 molecule A = ? atoms B<br>∴ 1 mol A = ? mol B | $\frac{1 \text{ mol A}}{? \text{ mol B}}$ | $\frac{? \text{ mol B}}{1 \text{ mol A}}$ |
| Example: $CO_2$ | 1 molecule $CO_2$ = 2 atoms O<br>∴ 1 mol $CO_2$ = 2 mol O | $\frac{1 \text{ mol } CO_2}{2 \text{ mol O}}$ | $\frac{2 \text{ mol O}}{1 \text{ mol } CO_2}$ |

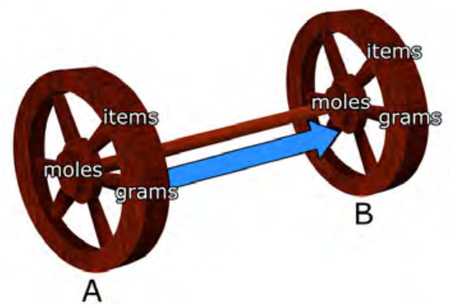
## Sample Problem — One-Step Conversion: Moles of A to Moles of B

How many moles of hydrogen are in 6.0 mol of water?

### What to Think about

1. Convert: mol $H_2O \rightarrow$ mol H
2. Setup: $6.0 \text{ mol } H_2O \times \dfrac{? \text{ mol H}}{1 \text{ mol } H_2O}$
3. Conversion factor:
   There are 2 H's in an $H_2O$;
   therefore there are
   2 mol of H in 1 mol of $H_2O$.
4. The values in the conversion factor (2 mol H per 1 mol $H_2O$) do not limit the significant figures in the answer as these values have no uncertainty. There are exactly 2 H's in an $H_2O$.

### How to Do It




$$6.0 \cancel{\text{mol } H_2O} \times \frac{2 \text{mol H}}{1 \cancel{\text{mol } H_2O}} = 12 \text{ mol H}$$

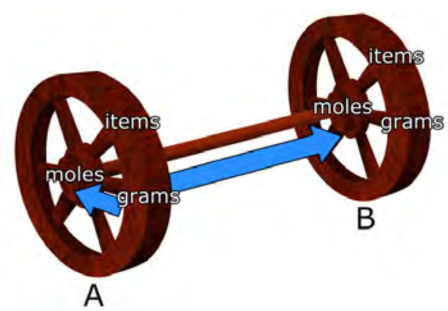
## Sample Problem — Two-Step Conversion: Mass of A to Moles of B

Hydrogen fuel cells are batteries that are continually supplied or "fuelled" with reactants. Their hydrogen sometimes comes from a process that "scrubs" it off methane molecules. How many moles of methane ($CH_4$) are required to get 0.860 g of hydrogen?

### What to Think about

1. Convert: g H $\rightarrow$ mol H $\rightarrow$ mol $CH_4$
2. Setup: $0.860 \text{ g H} \times \dfrac{1 \text{ mol H}}{? \text{ g H}} \times \dfrac{1 \text{ mol } CH_4}{? \text{ mol H}}$
3. Insert the conversion factors:
   1 mol H per 1.0 g H
   1 mol $CH_4$ per 4 mol H (given by the formula, $CH_4$)

### How to Do It




$$0.860 \cancel{\text{g H}} \times \frac{1 \cancel{\text{mol H}}}{1.0 \cancel{\text{g H}}} \times \frac{1 \text{ mol } CH_4}{4 \cancel{\text{mol H}}} = 0.22 \text{ mol } CH_4$$

## Sample Problem — Three-Step Conversion: Mass of A to Mass of B

How many grams of oxygen are in 14.6 g $CO_2$?

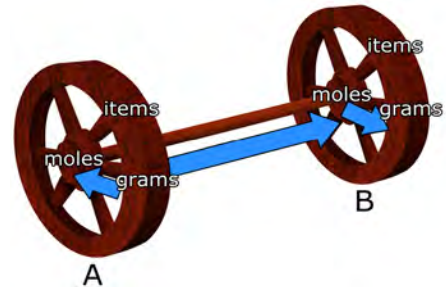
### What to Think about

1. Convert: g $CO_2 \rightarrow$ mol $CO_2 \rightarrow$ mol O $\rightarrow$ g O
2. Setup:

$$14.6 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{? \text{ g } CO_2} \times \frac{? \text{ mol O}}{1 \text{ mol } CO_2} \times \frac{? \text{ g O}}{1 \text{ mol O}}$$

3. Insert the conversion factors:
   1 mol $CO_2$ per 44.0 g $CO_2$
   2 mol O per 1 mol $CO_2$
   16.0 g O per 1 mol O

### How to Do It




$$14.6 \cancel{\text{ g } CO_2} \times \frac{1 \cancel{\text{ mol } CO_2}}{44.0 \cancel{\text{ g } CO_2}} \times \frac{2 \cancel{\text{ mol O}}}{1 \text{ mol } \cancel{CO_2}} \times \frac{16.0 \text{ g O}}{1 \cancel{\text{ mol O}}}$$

$= 10.6$ g O

## Practice Problems — One-, Two-, and Three-Step Conversions

1. Fill in the missing entries in these "axle" conversion factors.

   (a) $\dfrac{\_\_\_\_\_\_ \text{ mol O}}{\_\_\_\_\_\_ \text{ mol } SO_2}$ (b) $\dfrac{\_\_\_\_\_\_ \text{ mol } C_2H_4}{\_\_\_\_\_\_ \text{ mol H}}$

2. How many moles of $KNO_3$ contain 14 g of oxygen? ____________________

3. Fill in the missing entry in each conversion factor to determine how many oxygen atoms are in 2.5 g $K_2Cr_2O_7$.

$$2.5 \text{ g } K_2Cr_2O_7 \times \frac{1 \text{ mol } K_2Cr_2O_7}{\_\_\_\_\_\_ \text{ g } K_2Cr_2O_7} \times \frac{\_\_\_\_\_\_ \text{ mol O}}{1 \text{ mol } K_2Cr_2O_7} \times \frac{\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ \text{ atoms O}}{1 \text{ mol O}} =$$

4. What mass of sodium ions is contained in $1.23 \times 10^{24}$ formula units of sodium sulphide?

# 3.3 Activity: The Evaporation Rate of Water

### Question
What is the evaporation rate of water in molecules per second?

### Background
Evaporation occurs when a molecule on the surface of a liquid is struck with enough force by neighbouring molecules to break away from its attractions to those around it and enter the gas phase. Evaporation is often depicted like popcorn being popped. Pop…pop, pop, pop…pop, pop, etc. Let's use the mole concept to calculate the actual rate at which water molecules evaporate.

### Procedure
1. Half-fill a small beaker with water and weigh it. Record its mass and the time of day.
2. Put the beaker in a place where it won't be disturbed.
3. In about 30 min, weigh the beaker and its contents again, once more recording its mass and the time of day.

### Results and Discussion

| | Mass of Beaker and $H_2O$ (g) | Time of Day |
|---|---|---|
| **Initial** | | |
| **Final** | | |
| **Change** | | min |

1. Calculate the number of water molecules evaporated during the trial.

2. Calculate the duration of the trial in seconds.

3. Calculate the water's average evaporation rate during the trial in molecules per second.

4. Be suitably amazed. Oh come on, more amazed than that!

# 3.3 Review Questions

1. Acticoat dressings, developed in 1995 by Robert Burrell of the University of Alberta, are impregnated with crystals of silver that are only 15 nm (nanometres) in size. These nanocrystals are remarkably more effective at healing burns and other severe wounds than any treatment previously available. Acticoat bandages are credited with saving the lives and limbs of dozens of victims of the World Trade Center attack in New York City in 2001. What is the mass of a crystal containing $1.0 \times 10^3$ silver atoms?

2. Diamond is one way of arranging carbon atoms. The "Star of Africa" diamond, displayed with the crown jewels in the Tower of London, weighs 106.0 g and has an estimated value of over $400 million. How many carbon atoms compose the "Star of Africa" diamond?

3. What is the mass in grams of a chlorine atom?

4. How many propane molecules are in 72.6 g propane, $C_3H_8$?

5. On a particular day, 31.1 g (1 troy ounce) of gold cost $1300.
   (a) 31 g Au = _______________atoms of Au

   (b) How many atoms of gold could you buy for 1 cent on that day?

6. Complete the following "axle" conversion factors by filling in the appropriate numbers:

   (a) $\dfrac{\_\_\_\_\_\_ \text{ mol O}}{\_\_\_\_\_\_ \text{ mol } N_2O_4}$

   (b) $\dfrac{\_\_\_\_\_\_ \text{ mol } NO_2}{\_\_\_\_\_\_ \text{ mol N}}$

7. 2.3 mol $CO_2$ = _____________ mol O

8. Calcium oxalate is a poisonous compound found in rhubarb leaves. How many moles of carbon are in 52.4 mg of calcium oxalate?

9. Sodium phosphate is sold as a cleaner at most hardware stores. How many moles of sodium ions are there in $6.80 \times 10^{24}$ formula units of $Na_3PO_4$?

10. Sulphuric acid is used to produce a tremendous number and variety of materials including fertilizers, pigments, textiles, plastics, and explosives. What mass of sulphuric acid, $H_2SO_4$, contains 1.4 mol O?

11. How many carbon atoms are in 0.85 mol of the "pain-killer" acetaminophen, $C_8H_9NO_2$?

12. How many mercury(II) ions are in 100.0 g $HgCl_2$?

13. How many grams of chloride ions are in 8.3 g of copper(II) chloride?

14. What mass of carbon is present in $4.8 \times 10^{26}$ molecules of ethanol, $C_2H_5OH$?

15. Hydrogen fluoride, HF, can be used to etch glass. The white lines on the glassware in your lab may have been made by this acidic gas. Determine the mass in kilograms of $3.9 \times 10^{27}$ molecules of HF.

16. Up to $1.44 \times 10^5$ kg of various oxides of nitrogen are emitted by a gas-burning electrical plant in one year. Assuming this entire mass to be nitrogen dioxide, how many oxygen atoms would be present in this gas sample?

17. How many molecules are in 1.000 mg of the organic solvent, carbon tetrachloride?

18. Glycerol, $C_3H_5(OH)_3$, is a viscous, colourless liquid found in cough syrup, toothpaste, soaps, and many other household products. Calculate the number of hydrogen atoms in 4.5 mol of glycerol.

19. How many atoms are in 14.56 g of sodium hydrogen sulphate, the active ingredient in some toilet cleaners?

# 3.4 Molar Volume

## Warm Up

Recall from section 1.4 that the prefix "milli" means one-thousandth or $1 \times 10^{-3}$.

1. A milligram (mg) is one-_________________ of a gram.
2. A __________________ ( _______ ) is $1 \times 10^{-3}$ moles.
3. A millilitre (mL) is one-thousandth of a ______________.
4. 0.032 L = ___________ mL
5. 11.2 mg = ______________ g

## Molar Volume

Just as the mass of a mole of a substance is called its molar mass, the volume of a mole of a substance is called its molar volume. The **molar volume** of a substance is the space occupied by a mole of its particles. A solid's or a liquid's molar volume is determined by the size and spacing of its particles. The size of the particles has little effect on a gas's molar volume because the average distance between the particles is so much greater than their size.

Solids, liquids, and gases under constant pressure all expand when heated. Kinetic molecular theory explains that matter is composed of moving particles. At a higher temperature, a substance's particles are moving faster and are thereby hitting each other harder and bouncing farther apart. Since its particles have spread farther apart, a substance's molar volume is greater at higher temperatures.

Liquids and gases are more frequently measured by volume than by mass. A substance's molar volume allows you to convert the volume of the substance into its number of moles.

**Figure 3.4.1** *Gay-Lussac was an avid hot-air balloonist and conducted some of his experiments aloft.*

## Quick Check

1. What does the term "molar volume" mean? ________________________________________
2. A solid's or a liquid's molar volume is determined by the ___________ and _____________ of its particles.
3. A gas's molar volume is determined mainly by the ____________ of its particles.
4. What generally happens to the molar volume of a material as it is heated? _______________

## The Molar Volume of Gases

All of Dalton's evidence for the atomic theory came from combining mass ratios. During the same time period when Dalton lived, other scientists were following a separate line of research gathering data on combining volume ratios called volumetric data. In 1809, the French chemist, Joseph Gay-Lussac found that gases measured at the same temperature and pressure always reacted in whole-number volume ratios (Figure 3.4.1). For example, two volumes of hydrogen gas and one volume of oxygen gas react to produce two volumes of gaseous water.

Using Gay Lussac's findings, the Italian chemist Amedeo Avogadro hypothesized that equal volumes of different gases, measured at the same temperature and pressure, have equal numbers of particles. Modern chemists still refer to this as **Avogadro's hypothesis**.

At low pressures, the different sizes and attractive forces of different particles have little effect on the gas's volume because the particles are so far apart on average. As an example, 1 mol of any gas at 0°C and 101.3 kPa occupies approximately 22.4 L. Chemists refer to 0°C and 101.3 kPa as **standard temperature and pressure** or **STP** for short. The standard pressure 101.3 kPa is normal atmospheric pressure at sea level.

The molar volume of any gas at STP is approximately 22.4 L.

| Name | Equivalence Statement | Conversion Factors | |
|---|---|---|---|
| Molar gas volume | 1 mol = 22.4 L @ STP | $\frac{22.4\ L}{1\ mol}$ @ STP | $\frac{1\ mol}{22.4\ L}$ @ STP |

While all gases at the same temperature and pressure have approximately the same molar volume, each solid and liquid has its own characteristic molar volume. In questions requiring volumetric conversions, you will be given either the molar volume of any solid, liquid, or gas not at STP, or you will be given a means to calculate it.

## Sample Problem — Converting Moles to Volume

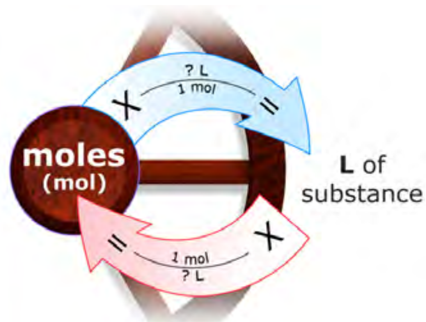
Atmospheric nitrogen and oxygen react during lightning storms to produce nitrogen monoxide that is quickly converted to nitrogen dioxide. What is the volume of 1.3 mol of $NO_2$ at STP?

### What to Think about

1. Convert: mol $NO_2 \rightarrow$ L $NO_2$
2. Setup:
   $1.3\ mol\ NO_2 \times \frac{?\ L\ NO_2}{1\ mol\ NO_2}$
3. Conversion factor:
   22.4 L $NO_2$ per 1 mol $NO_2$

### How to Do It

$$1.3\ \cancel{mol\ NO_2} \times \frac{22.4\ L\ NO_2}{1\ \cancel{mol\ NO_2}} = 29\ L\ NO_2$$




## Sample Problem — Converting Volume to Moles

600.0 L of air at STP is compressed into a scuba tank. How many moles of air are in the tank?

### What to Think about

1. Convert: L air $\rightarrow$ mol air
2. Setup: $600.0\ L\ air \times \frac{1\ mol\ air}{?\ L\ air}$
3. Conversion factor:
   1 mol air per 22.4 L air

### How to Do It

$$600.0\ \cancel{L\ air} \times \frac{1\ mol\ air}{22.4\ \cancel{L\ air}} = 26.8\ mol\ air$$

### Practice Problems — Converting Moles to Volume and Volume to Moles

1. What volume of oxygen gas at STP contains 1.33 mol of $O_2$?

2. In British Columbia, the burnt-match odor of sulphur dioxide is often associated with pulp and paper mills. How many moles of $SO_2$ are in 9.5 L of $SO_2$ at STP?

3. Silicon dioxide, better known as quartz, has a molar volume of 22.8 $cm^3$/mol. What is the volume of 0.39 mol of $SiO_2$?

## Multi-Step Conversions Involving the Volume of a Substance

Recall from section 3.3 that as you learn more chemical quantities we'll add them to the rim of the wheel. You'll relate or connect the volume of a substance to the other quantities through the mole. For example, if you wanted to convert the volume of a substance into its mass, you would convert litres into moles and then moles into grams.

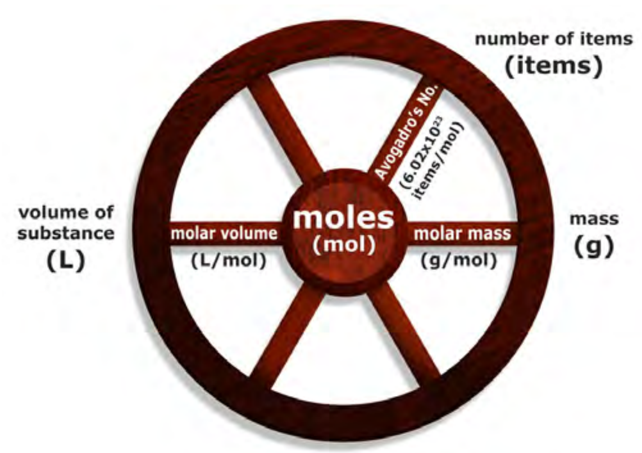



**Figure 3.4.2** *Use the wheel model to help you do conversions involving the volume of a substance.*

### Sample Problem — Two Step Conversion: Volume to Number of Items (Atoms)

The gas in neon signs is at extremely low pressure. How many neon atoms are present in a sign containing 75 mL of neon gas at a molar volume that is 100 times greater than the molar volume at STP?

#### What to Think about

1. Convert: mL → L
2. Convert: L Ne → mol Ne → atoms Ne
3. Setup:

$$0.075 \text{ L Ne} \times \frac{1 \text{ mol Ne}}{? \text{ L Ne}} \times \frac{? \text{ atoms Ne}}{1 \text{ mol Ne}}$$

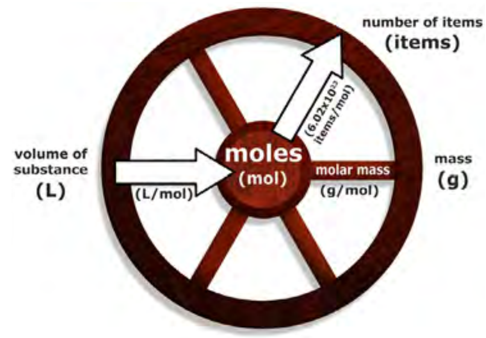
4. Conversion factors:
   2240 L Ne per 1 mol Ne
   $6.02 \times 10^{23}$ atoms Ne per 1 mol Ne

#### How to Do It

$$75 \text{ ml} \times \frac{1.0 \text{ L}}{1000 \text{ mL}} = 0.075 \text{ L}$$

$$0.075 \cancel{\text{ L Ne}} \times \frac{1 \cancel{\text{ mol Ne}}}{2240 \cancel{\text{ L Ne}}} \times \frac{6.02 \times 10^{23} \text{ atoms Ne}}{1 \cancel{\text{ mol Ne}}}$$

$$= 2.0 \times 10^{19} \text{ atoms Ne}$$




### Sample Problem — Two Step Conversion: Volume to Mass

Natural gas is used to heat many homes and may fuel the Bunsen burners in your laboratory. Natural gas consists primarily of methane, $CH_4$. What is the mass of 8.0 L of $CH_4$ at STP?

#### What to Think about

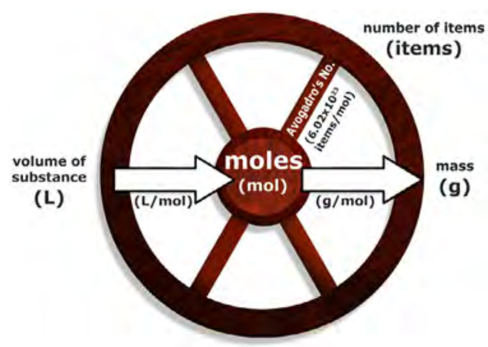
1. Convert: L $CH_4$ → mol $CH_4$ → g $CH_4$
2. Setup:

$$8.0 \text{ L } CH_4 \times \frac{1 \text{ mol } CH_4}{? \text{ L } CH_4} \times \frac{? \text{ g } CH_4}{1 \text{ mol } CH_4}$$

3. Conversion factors:
   22.4 L $CH_4$ per 1 mol $CH_4$
   16.0 g $CH_4$ per 1 mol $CH_4$

#### How to Do It

$$8.0 \text{ L } CH_4 \times \frac{1 \cancel{\text{ mol } CH_4}}{22.4 \cancel{\text{ L } CH_4}} \times \frac{16.0 \cancel{\text{ g } CH_4}}{1 \text{ mol } \cancel{CH_4}} = 5.7 \text{ g } CH_4$$



### Sample Problem — Three Step Conversion: Mass of A to Volume of B

People often refer to the amount of $CO_2$ produced as the carbon "footprint" of a process. What volume of $CO_2$ at STP contains 0.20 g of carbon?

#### What to Think about

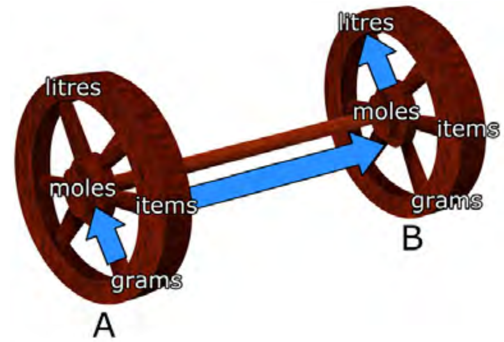
1. Convert:
   g C $\rightarrow$ mol C $\rightarrow$ mol $CO_2$ $\rightarrow$ L $CO_2$

2. Setup:
   $$0.20\text{ g C} \times \frac{1\text{ mol C}}{?\text{ g C}} \times \frac{1\text{ mol CO}_2}{?\text{ mol C}} \times \frac{?\text{ L CO}_2}{1\text{ mol CO}_2}$$

3. Conversion factors:
   1 mol C per 12.0 g C
   1 mol $CO_2$ per 1 mol C
   22.4 L $CO_2$ per 1 mol $CO_2$

#### How to Do It

$$0.20\ \cancel{\text{g C}} \times \frac{1\ \cancel{\text{mol C}}}{12.0\ \cancel{\text{g C}}} \times \frac{1\ \cancel{\text{mol CO}_2}}{1\ \cancel{\text{mol C}}} \times \frac{22.4\text{ L CO}_2}{1\ \cancel{\text{mol CO}_2}}$$
$$= 0.37\text{ L CO}_2$$




*Converting mass of one substance to volume of another*

### Practice Problems — Conversions: Volume to Number of Items or Mass; Mass to Volume

1. $H_2S$ gas is released from rotten eggs. What volume of $H_2S$ gas at STP contains 17 g $H_2S$?

2. Fill in the missing entry in each conversion factor below to determine the mass of carbon in 1.0 L of propane, $C_3H_8$, at STP.
   $$1.0\text{ L C}_3\text{H}_8 \times \frac{1\text{ mol C}_3\text{H}_8}{\_\_\_\_\_\text{ L C}_3\text{H}_8} \times \frac{\_\_\_\_\_\text{ mol C}}{1\text{ mol C}_3\text{H}_8} \times \frac{\_\_\_\_\_\text{ g C}}{1\text{ mol C}}$$

3. Ethylene glycol ($C_2H_6O_2$) is widely used as an automotive antifreeze. The molar volume of $C_2H_6O_2$ is 0.0559 L/mol. How many hydrogen atoms are in 200.0 mL of $C_2H_6O_2$?

## Molar Volume and Density

**Density** is the amount of matter in a given volume of an object or material. In other words, it is the mass per unit volume. Density is a conversion factor that relates a substance's mass directly to its volume without any reference to the mole. In terms of our wheel model, density is the section of the rim between grams and litres.

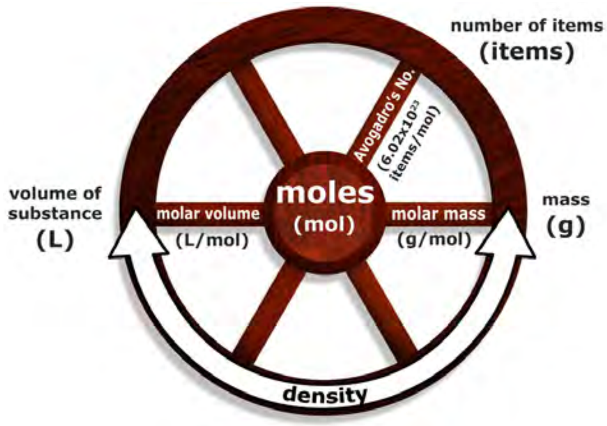



**Figure 3.4.3** *Interconverting the mass and the volume of a substance*

### Sample Problem — Converting Volume Directly to Mass

The density of methanol, $CH_3OH$, at 20°C is 0.813 g/mL. What is the mass of 0.500 L of the alcohol at 20°C?

| What to Think about | How to Do It |
| --- | --- |
| 1. Convert: L $CH_3OH \rightarrow$ g $CH_3OH$<br>2. Setup: $0.500\text{ L }CH_3OH \times \frac{?\text{ g }CH_3OH}{1\text{ L }CH_3OH}$<br>3. Conversion factor: 813 g/L | $0.500\ \cancel{\text{L }CH_3OH} \times \frac{813\text{ g }CH_3OH}{1\ \cancel{\text{L }CH_3OH}} = 407\text{ g }CH_3OH$ |

It is the densities of substances, rather than their molar volumes, that are usually published in reference texts and tables. Molar volume and density are related through molar mass. An examination of their units (dimensional analysis) reveals that:

$$\frac{\cancel{g}}{\text{mol}} \times \frac{\text{L}}{\cancel{g}} = \frac{\text{L}}{\text{mol}}$$

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

## Sample Problem — Calculating Molar Volume from Density

In an episode of the television show "MythBusters," the team floated an aluminum foil boat on the invisible gas, sulphur hexafluoride, $SF_6$. $SF_6$ has a density of 6.00 g/L at room temperature and pressure, about six times that of air. What is the molar volume of $SF_6$ under these conditions?

| What to Think about | How to Do It |
|---|---|
| 1. molar volume = $\frac{\text{molar mass}}{\text{density}}$<br>2. Setup: $\frac{?\text{ g } SF_6}{1\text{ mol } SF_6} \times \frac{1\text{ L } SF_6}{6.00\text{ g } SF_6}$<br>3. Conversion factors:<br>146.1 g $SF_6$ per 1 mol $SF_6$ | $\frac{146.1\ \cancel{\text{g } SF_6}}{1\text{ mol } SF_6} \times \frac{1\text{ L } SF_6}{6.00\ \cancel{\text{g } SF_6}} = 24.3\ \frac{\text{L } SF_6}{\text{mol } SF_6}$ |

## Practice Problems — Calculating Molar Volume and Density

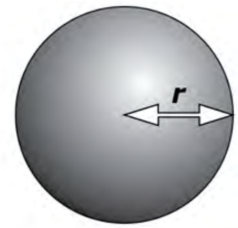
1. Gold has a density of 19.42 g/cm$^3$. The standard gold bar held as gold reserves by central banks weighs 12.4 kg. What is the volume of the standard gold bar?

2. Mercury has a density of 13.534 g/mL at room temperature. What is the mass of 12.7 mL of mercury?

3. Although ethanol ($C_2H_5OH$) is best known as the type of alcohol found in alcoholic beverages, its largest use is as a fuel or fuel additive. The density of ethanol is 0.789 g/mL. What is the molar volume of ethanol?

# 3.4 Activity: The Atomic Radius of Aluminum




## Question

What is the radius of an aluminum atom?

## Background

We can't even see an individual atom with the naked eye but we can still derive an atom's radius from the molar volume and packing arrangement of a substance's atoms. Molar volume is the actual amount of space required to "house" 1 mol of the atoms and includes the space between them. In a technique called X-ray diffraction, chemists reflect X-rays off the substance. The scattering (diffraction) pattern allows chemists to determine how the atoms are arranged. Aluminum atoms are packed in such a way that 74% of the metal's volume is the volume of the atoms themselves and the rest is space.

## Procedure

1. Weigh and measure the dimensions of a small aluminum block and fill in the table below.

| Mass (g) | Length (cm) | Width (cm) | Height (cm) | Volume ($cm^3$) | Density ($g/cm^3$) |
|---|---|---|---|---|---|
| | | | | | |

## Results and Discussion

1. Calculate the molar volume of aluminum in $cm^3/mol$.

2. Calculate the volume of 1 mol of Al atoms (excluding the space between them).

   Volume = 0.74 × ______________ $cm^3/mol$ = ________________ $cm^3/mol$

3. Calculate the volume of one Al atom.

$$\text{Atomic volume} = \frac{\_\_\_\_\_\_\ \text{cm}^3\ \text{Al}}{1\ \text{mol Al}} \times \frac{1\ \text{mol Al}}{\_\_\_\_\_\_\_\_\_\_\ \text{atoms Al}} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\ \text{cm}^3/\text{mol}$$

4. Calculate the radius of a spherical aluminum atom. Given $V = {}^4/_3\, \pi\, r^3$ solve for $r$.

5. Convert the radius of the atoms into nanometres.

6. The accepted value for the radius of an aluminum atom is 0.143 nm. What is your percent error?

You just figured out the radius of an atom. How amazing is that! Cu (8.96 $g/cm^3$) atoms are packed the same way as Al atoms if you'd like to repeat this activity for copper.

# 3.4 Review Questions

1. Liquid octane, $C_8H_{18}$, has a molar volume of 82.4 mL/mol. What is the volume of 250 millimoles of $C_8H_{18}$?

2. How many moles of air are there in a human lung with a volume of 2.4 L at STP?

3. 2.75 L $N_2$ at STP = ______________ mol $N_2$

4. Air is approximately 21% oxygen. How many moles of oxygen are in 5.0 L of air at STP?

5. Diphosphorus pentoxide is a gas produced each time you strike a match. What is the mass of 2.57 L of this gas at STP?

6. A 525 mL flask contains 0.935 g of a noble gas at STP. Identify the gas from its molar mass.

7. Acetylene gas, $C_2H_2$, is used as a fuel in welding torches. How many acetylene molecules are in a cylinder that delivers 1400 L of acetylene at STP?

8. $5 \times 10^{19}$ molecules $PH_3$ = ___________ mL $PH_3$ at STP

9. Propane gas, $C_3H_8(g)$, is easily compressible to a storable liquid. A standard barbecue tank holds 9.1 kg of propane. How many litres of gas will the tank release at STP?

10. Soft drinks are bottled under pressure forcing $CO_2$ into solution. The industry expresses the amount of carbonation in volumes of $CO_2$ at STP per volume of solution. The carbonation of a typical soft drink is 3.7 v/v meaning that a 355 mL can contains $3.7 \times 355$ mL $CO_2$ at STP. What is the mass of $CO_2$ in a 355 mL can?

11. How many moles of hydrogen are in 83.9 L of ammonia gas, $NH_3$, at STP?

12. Nitrous oxide, $N_2O$, is commonly called "laughing gas." It is sometimes used by dentists as a partial anaesthetic. How many grams of nitrogen are in 3.84 L of $N_2O$ at STP?

13. Dinitrogen tetroxide is one of the most important rocket propellants ever developed. How many oxygen atoms are in 27.2 L of the gas at STP?

14. Disposable lighters often contain butane, $C_4H_{10}$ (density = 0.601 g/mL). How many grams of butane are there in a lighter containing 15 mL of the fuel?

15. Mercury is a liquid metal with a density of 13.546 g/mL at 20°C. What is the molar volume of mercury at 20°C?

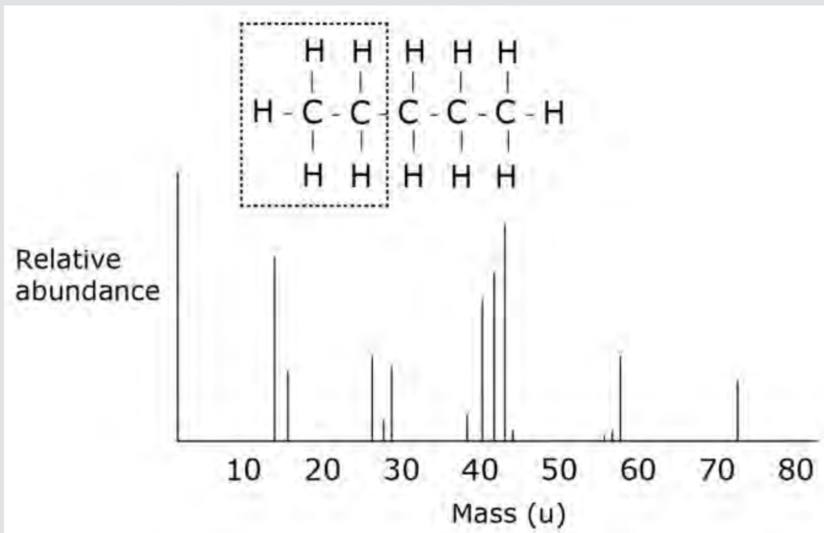
16. Gold has a density of 19.42 g/cm$^3$. How many moles of gold are there in a 5.0 cm$^3$ strip?

17. Liquid bromine, $Br_2$, has a density of 3.53 g/mL. How many bromine molecules are in 15.0 mL of bromine?

# 3.5 Composition Analysis — Determining Formulas

## Warm Up

Forensic investigators collect samples from crime scenes. How do technicians identify the unknown samples? An instrument called a **mass spectrometer** can identify the vast majority of compounds. Each compound has a unique mass spectrum; much like each person has a unique fingerprint. A mass spectrometer breaks most of the molecules into fragments. In so doing, it creates a variety of particles from individual atoms to the intact molecule itself, and then marks the mass of each of these particles along a graph's horizontal axis. The height of the line in the spectrum indicates the relative abundance of that particle. Below is a simplified mass spectrum of a compound called pentane ($C_5H_{12}$).




*Mass spectrum analysis of pentane*

1. The last spectral line represents the intact molecule. What is its molecular mass? __________
2. Draw an arrow to point to the spectral line that represents the mass of the outlined fragment.
3. Why do you think there are more of some fragments than others?

______________________________________________________________________________

______________________________________________________________________________

The important point here is that one of the most sophisticated tools in a chemist's arsenal identifies compounds solely from their "mass profile". In this section, you'll learn to determine a compound's formula from its composition by mass.

### Percentage Composition

**Percentage composition** is the percent of a compound's mass contributed by each type of atom in the compound.

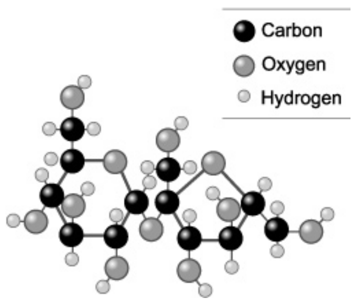
A compound's percentage composition can be determined theoretically from its formula.

## Sample Problem — Determining Percentage Composition

What is the percentage composition of a sugar with the formula $C_{12}H_{22}O_{11}$?

### What to Think about

1. Calculate the sugar's molar mass.
2. Thus one mole of this sugar contains 144 g C, 22 g H, and 176 g O.
3. Express each element's percentage of the molar mass.




*A sugar molecule with 12 carbon atoms, 22 hydrogen atoms, and 11 oxygen atoms.*

### How to Do It

$$\frac{12 \text{ mol C}}{1 \text{ mol } C_{12}H_{22}O_{11}} \times \frac{12.0 \text{ g C}}{1 \text{ mol C}} = \frac{144.0 \text{ g C}}{1 \text{ mol } C_{12}H_{22}O_{11}}$$

$$\frac{22 \text{ mol H}}{1 \text{ mol } C_{12}H_{22}O_{11}} \times \frac{1.0 \text{ g H}}{1 \text{ mol H}} = \frac{22.0 \text{ g H}}{1 \text{ mol } C_{12}H_{22}O_{11}}$$

$$\frac{11 \text{ mol O}}{1 \text{ mol } C_{12}H_{22}O_{11}} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}} = \frac{176.0 \text{ g O}}{1 \text{ mol } C_{12}H_{22}O_{11}}$$

$$\text{Total} = 342.0 \text{ g/mol } C_{12}H_{22}O_{11}$$

$$\% \text{ C} = \frac{144.0 \text{ g C}}{342.0 \text{ g } C_{12}H_{22}O_{11}} \times 100 = 42.1\ \%$$

$$\% \text{ H} = \frac{22.0 \text{ g H}}{342.0 \text{ g } C_{12}H_{22}O_{11}} \times 100 = 6.4\ \%$$

$$\% \text{ O} = \frac{176.0 \text{ g O}}{342.0 \text{ g } C_{12}H_{22}O_{11}} \times 100 = 51.5\ \%$$

## Practice Problems — Determining Percentage Composition

1. Ibuprofen is a common pain reliever and anti-inflammatory. Its formula is $C_{13}H_{18}O_2$. What is its percentage composition?

2. Ammonium sulphate, $(NH_4)_2SO_4$, is a common fertilizer used to lower the pH of soil. Calculate its percentage composition.

3. Many salts are hydrated, which means they have water molecules incorporated into their ionic crystal lattice in a fixed ratio. Magnesium sulphate heptahydrate, $MgSO_4 \cdot 7\ H_2O$, has seven water molecules incorporated into the crystal lattice for each magnesium ion and sulphate ion. Calculate the percentage of water by mass in $MgSO_4 \cdot 7\ H_2O$.

## Empirical, Molecular, and Structural Formulas

Every molecular compound has three formulas; an empirical formula, a molecular formula, and a structural formula. They become more specific in that order.

- The **empirical formula** is the simplest integral ratio of the different types of atoms in the compound.
- The **molecular formula** is the actual number of each type of atom in each molecule of the compound.
- The **structural formula** shows how the atoms in a molecule are arranged. It is a diagram that shows the pattern of the atoms' connections.

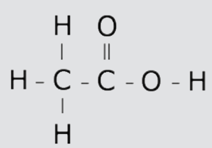
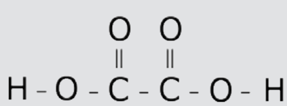
Organic chemistry is the study of compounds and reactions involving carbon. There are millions of organic compounds. Glucose is an organic compound with a molecular formula of $C_6H_{12}O_6$. The subscripts 6, 12, 6 can be reduced or simplified to 1, 2, 1. We don't show the number 1 as a subscript in a formula so the empirical formula of glucose is $CH_2O$.

Many compounds have the same empirical formula but different molecular formulas. Their molecular formulas all reduce to the same ratio. For example, all alkenes such as ethene ($C_2H_4$), propene ($C_3H_6$), and butene ($C_4H_8$), have an empirical formula of $CH_2$ because each of their molecular formulas can be reduced to a 1 to 2 ratio.

Compounds with the same molecular formula but with different structural formulas, i.e. the same atoms are put together differently, are called structural isomers. For example, $C_4H_{10}$ has two structural isomers. You will learn more about structural isomers in chapter 8.

### Quick Check

1. Complete the following table.

| Structural Formula | Molecular Formula | Empirical Formula |
|---|---|---|
|  | | |
|  | | |

## Determining an Empirical Formula from Percent Composition

In section 3.1 you learned how to determine an element's relative atomic mass from a compound's percent composition and formula. Earlier in this section (3.5), you learned how to determine the percent composition of a compound from its formula and the atomic masses of its component elements. There's only one more arrangement of these variables to learn. That's how to determine the formula of a compound from its percent composition and the atomic masses of its component elements.

The word "empirical" is an adjective meaning that something is based on observation or experiment. Empirical formulas are determined from the mass ratios of a compound's component elements; in other words, from its percent composition. The most direct, but not always easiest, way to experimentally determine a compound's percent composition is to decompose a sample of the compound into its component elements.

## Sample Problem — Determining an Empirical Formula

Determine the empirical formula of a compound that is 48.65% carbon, 8.11% hydrogen, and 43.24% oxygen.

### What to Think about

1. In 100.0 g of the substance, there would be 48.65 g C, 8.11 g H, and 43.24 g O. Convert these amounts into moles.
2. Divide each molar quantity by the smallest one and then multiply by whatever factor is necessary to find their integral ratio (as shown in a conventional formula).

   The mole ratio and the individual atom ratio are of course the same. This means the subscripts in a formula can be read either as mole ratios or as individual atom ratios. If this compound has 3 mol of carbon atoms for every 2 mol of oxygen atoms then it has 3 dozen carbon atoms for every 2 dozen oxygen atoms, and 3 carbon atoms for every 2 oxygen atoms.

### How to Do It

$$48.65 \cancel{\text{g C}} \times \frac{1 \text{ mol C}}{12.0 \cancel{\text{g C}}} = 4.0542 \text{ mol C}$$

$$8.11 \cancel{\text{g H}} \times \frac{1 \text{ mol H}}{1.0 \cancel{\text{g H}}} = 8.1100 \text{ mol H}$$

$$43.24 \cancel{\text{g O}} \times \frac{1 \text{ mol O}}{16.0 \cancel{\text{g O}}} = 2.7025 \text{ mol O}$$

$$\frac{C_{4.0542}H_{8.1100}O_{2.7025}}{2.7025} = C_{1.5}H_3O$$

$$2\,(C_{1.5}H_3O) = C_3H_6O_2$$

## Practice Problems — Determining an Empirical Formula

1. A compound is 18.7% Li, 16.3% C, and 65.5% O. Determine its empirical formula.

2. A compound is 9.93% C, 58.6% Cl, and 31.4% F. Determine its empirical formula.

3. A sample of a compound contains 5.723 g Ag, 0.852 g S, and 1.695 g O. Determine its empirical formula.

## Determining the Molecular Formula of a Compound

Recall the difference between the empirical formula and the molecular formula of a compound. The empirical formula is the simplest integral ratio of the different types of atoms in the compound. The molecular formula is the actual number of each type of atom in each molecule. A compound's molecular formula is an integral multiple of its empirical formula. Its molecular formula's molar mass is that same integral multiple of its empirical formula's molar mass. For example, butane has an empirical formula of $C_2H_5$
(29 g/mol) and a molecular formula of $C_4H_{10}$ (58 g/mol). The "molecular formula's molar mass" is just another way of saying "the compound's molar mass." Therefore, we can determine the compound's molecular formula from its empirical formula and its molar mass.

$$\text{molecular formula} = \text{empirical formula} \times \frac{\text{compound's molar mass}}{\text{molar mass of empirical formula}}$$

There are many ways to experimentally derive a compound's molar mass. If the compound is volatile, meaning it is easily evaporated, then you know from section 3.4 that 1 mol of any gas occupies 22.4 L at STP. The mass of 22.4 L at STP thus provides the compound's molar mass.

### Sample Problem — Determining a Molecular Formula

A compound has an empirical formula of $CH_2$ and a molar mass of 42.0 g/mol. Determine its molecular formula.

| What to Think about | How to Do It |
|---|---|
| 1. Calculate the molar mass of the empirical formula. | $\frac{1\text{ mol C}}{1\text{ mol }CH_2} \times \frac{12.0\text{ g C}}{1\text{ mol C}} = \frac{12.0\text{ g C}}{1\text{ mol }CH_2}$ <br> $\frac{2\text{ mol H}}{1\text{ mol }CH_2} \times \frac{1.0\text{ g H}}{1\text{ mol H}} = \frac{2.0\text{ g H}}{1\text{ mol }CH_2}$ <br> Total $= 14.0\text{ g/mol }CH_2$ |
| 2. Divide the molar mass of the molecular formula by the molar mass of the empirical formula. | $\frac{42.0\text{ g/mol}}{14.0\text{ g/mol}} = 3.00$ |
| 3. Multiply the empirical formula itself by this factor. | $3\ CH_2 = C_3H_6$ |

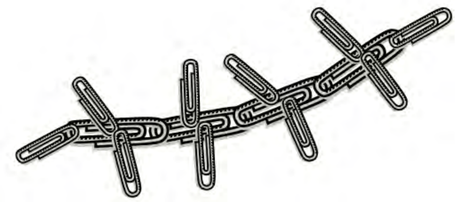
## Practice Problems — Determining a Molecular Formula

1. Vinegar is a dilute solution of acetic acid. The molar mass of acetic acid is 60.0 g/mol and it has an empirical formula of $CH_2O$. What is the molecular formula of acetic acid?

2. A compound has an empirical formula of $C_3H_4$. Which of the following are possible molar masses of the compound: 20 g/mol, 55 g/mol, 80 g/mol, 120 g/mol?

3. A small sample of antifreeze was analyzed. It contained 4.51 g C, 1.13 g H, and 6.01 g O. From the elevation of water's boiling point, it was determined that the antifreeze's molar mass is 62.0 g/mol. What is the molecular formula of antifreeze?

# 3.5 Activity: Determining the Empirical Formula of Butane from the Percentage Composition of Its Model



## Question
What is the empirical formula of butane?

## Background
Recall that "empirical" is an adjective meaning "based on observation or experiment." Empirical formulas are determined from the mass ratios of a compound's component elements (i.e., from its percent composition as determined by analysis). The empirical formula is the simplest integral ratio of the different types of atoms in the compound.

## Procedure
1. Use #1 (regular sized) paper clips to represent hydrogen atoms. Jumbo paper clips represent carbon atoms. As a prelude to this exercise, someone must weigh 48 of each type of paperclip and divide by 4 to obtain the mass per dozen. Provide these values to the students to enter in column 3.
2. Form a group of two to five students.
3. Each student links 4 jumbo paper clips together with 10 regular paper clips.
4. Unlink all the clips and weigh all your group's jumbo clips together. Record the mass of your group's jumbo clips as the mass of carbon in the table below.
5. Weigh all your group's regular sized clips together. Record the mass of your group's regular clips as the mass of hydrogen in the table below.

## Results and Discussion

| Element | Mass (g) | Mass per Dozen (g/doz) | Number (doz) | Dozen Ratio | Empirical Formula |
|---|---|---|---|---|---|
| carbon | | | | 1.0 | |
| hydrogen | | | | | |

1. Calculate how many dozens of each type of paper clip are in your group's sample.

2. Calculate the dozen ratio to find out how many dozens of hydrogen atoms there are for each dozen carbon atoms.

3. By what integer do you need to multiply this ratio in order to obtain an integral dozen ratio? ______________

4. What is the empirical formula of butane? __________________________________________________

5. Given the molecular models you made, what is the molecular formula of butane?

## 3.5 Review Questions

1. Menthol is a strong-smelling compound that is used in cough drops. It has a formula of $C_{10}H_{20}O$. Calculate its percentage composition.

2. Sodium acetate trihydrate ($NaCH_3COO \cdot 3H_2O$) is a salt commonly used in pickling foods. Calculate the percentage of water by mass in this compound.

3. Trinitrotoluene ($C_7H_5O_6N_3$) is an explosive commonly referred to as TNT. Calculate the percentage of nitrogen by mass in this compound.

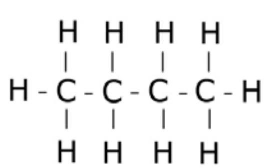
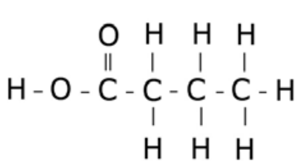
4. Complete the following table.

5. (a) Explain why the empirical formula alone is not enough to identify a compound.

   (b) What other piece of information will allow you to determine its molecular formula?

6. A pigment on a suspected forgery is analyzed using X-ray fluorescence and found to contain 0.5068 mol Ba, 0.5075 mol C, and 1.520 mol O. Determine its empirical formula.

7. A sample of caffeine is analyzed and found to contain 1.4844 g C, 0.1545 g H, 0.4947 g O and 0.8661 g N. Determine the empirical formula of caffeine.

| Structural Formula | Molecular Formula | Empirical Formula |
|---|---|---|
|  | | |
|  | | |

8. (a) In a TV series, a forensic anthropologist uses X-ray fluorescence to analyze a dental filling found in skeletal remains. The results of the analysis are provided as *atomic* percentages: 2.85% Al, 87.4% Si, and 9.75% Yb. Convert these results into mass percentages.

   (b) These results identified the filling as a commercial restorative material called Heliomolar. How might identifying the material be useful in helping to identify the remains?

9. A compound has an empirical formula of $NH_2$ and a molar mass of 32.1 g/mol. What is the compound's molecular formula?

10. A sample of ascorbic acid, also known as vitamin C, was analyzed and found to contain 1.080 g C, 0.121 g H, and 1.439 g O. Ascorbic acid has a molar mass of 176.1 g/mol. Determine the molecular formula of ascorbic acid.

11. A hydrocarbon is a compound containing only carbon and hydrogen. One particular hydrocarbon is 92.29% carbon by mass. If the compound's molar mass is 78.0 g/mol then what is its molecular formula?

12. Cannizzaro determined that a certain compound of carbon and oxygen had a molecular mass of 44.0 u. This meant that a certain volume of this gaseous compound weighed 44.0 times as much as the same volume of hydrogen gas at the same temperature and pressure. This compound was analyzed and found to be 27.3% carbon by mass.

    (a) What is the total mass of carbon in a molecule of this compound?

    (b) Cannizzaro repeated this experiment on many carbon compounds. Because he never found a molecule with less carbon than this one, Cannizzaro assumed that this molecule had only one carbon atom. Was this assumption correct?

# 3.6 Molar Concentration

*Examples of common household liquids*

## Warm Up

1. List three products in your refrigerator that are solutions.
______________________________________________
2. Name some substances that are dissolved in these solutions.
______________________________________________
3. Where else in your home are solutions kept?
______________________________________________

### Molarity — A Useful Unit of Concentration

Many chemicals are dispensed in solution and most chemical reactions occur in solution.Recall from chapter 2 that a solution is a type of mixture in which the chemical species are completely mixed. A solute is a minor component of the mixture, generally what has been dissolved. The solvent is the major component of the mixture, generally what the solute was dissolved in.

**Concentration** is any expression of the proportion of a chemical in a solution. Chemists need to know the amount of solute present in any volume of solution they might dispense. Therefore, concentration is most usefully expressed as an amount of solute per volume of solution rather than per volume of solvent. There are many units of concentration. Common units of concentration express the amount of solute in grams. These include grams per litre of solution, percent m/v, which is the number of grams (mass) per 100 mL (volume) of solution, and parts per million (ppm), when expressed as the number of grams per million grams of solution. A useful unit of concentration for chemists expresses the quantity of solute in moles.

**Molarity (M)** is the number of moles of the chemical per litre of solution.

For example, 1.8 M HCl means 1.8 mol HCl per litre of solution. Molar concentrations allow chemists to directly compare the number of particles in the same volume of different solutions. For example, 10 mL of 2 M $Li^+$ contains twice as many ions as 10 mL of 1 M $Na^+$.

| Name | Equivalence Statement | Conversion Factors | |
|---|---|---|---|
| Molar concentration | 1 L solution = ? mol solute | $\frac{\text{? mol solute}}{\text{1 L solution}}$ | $\frac{\text{1 L solution}}{\text{? mol solute}}$ |
| Example: 3 M HCN | 1 L solution = 3 mol HCN | $\frac{\text{3 mol HCN}}{\text{1 L solution}}$ | $\frac{\text{1 L solution}}{\text{3 mol HCN}}$ |

## Quick Check

1. Give one reason why solutions are important in chemistry.

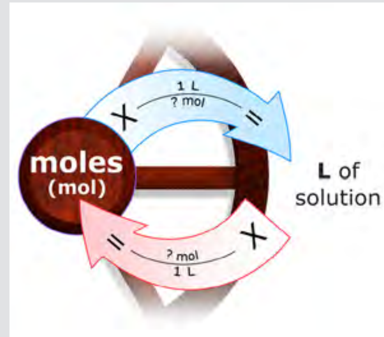
   ______________________________________________

2. What does 2 M NaOH mean?

   ______________________________________________

3. Why is molarity a useful unit of concentration?

   ______________________________________________




## Sample Problem — Converting Volume of Solution into Moles of Solute

The average salinity (total salt concentration) of seawater is 0.60 M. How many moles of salt are in a toy bucket containing 435 mL of seawater?

### What to Think about

1. Convert: mL → L
2. Convert: L soln → mol salt
3. Setup: $0.435 \text{ L soln} \times \dfrac{? \text{ mol salt}}{1 \text{ L soln}}$
4. Conversion factor:
   0.60 mol salt per 1 L soln

### How to Do It

$$435 \text{ mL} \times \frac{1.0 \text{ L}}{1000 \text{ mL}} = 0.435 \text{ L}$$

$$0.435 \cancel{\text{L soln}} \times \frac{0.60 \text{ mol salt}}{1 \cancel{\text{L soln}}} = 0.26 \text{ mol salt}$$

Note: There is no conventional abbreviation for "solution" but we will use "soln" in our calculations.

## Sample Problem — Converting Moles of Solute into Volume of Solution

What volume of 3.0 M HCl should a chemist dispense to obtain 0.25 mol HCl?

### What to Think about

1. Convert: mol HCl → L soln
2. Setup: $0.25 \text{ mol HCl} \times \dfrac{1 \text{ L soln}}{? \text{ mol HCl}}$
3. Conversion factor:
   1 L soln per 3.0 mol HCl

### How to Do It

$$0.25 \cancel{\text{mol HCl}} \times \frac{1 \text{ L soln}}{3.0 \cancel{\text{mol HCl}}} = 0.083 \text{ L soln}$$

### Practice Problems — Converting Moles of Solute to/from Volume of Solution

1. 0.72 L of 2.5 M NaOH = __________ mol NaOH

2. An intravenous bag of saline solution contains 0.154 M NaCl. How many moles of NaCl does a 500.0 mL bag contain?

3. 3.0 mol HCl = __________L of 0.60 M HCl

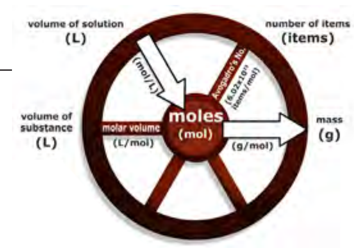
4. A person's urine may have a distinct odor as soon as 15 min after eating asparagus. Methanethiol, one of the metabolic products responsible for this odor, can be detected by some people in concentrations as low as $4.0 \times 10^{-8}$ M. At this concentration, what volume of urine contains 1.0 mmol of methanethiol?

## Preparing a Standard Solution from a Solid

A solution of known concentration is called a **standard solution**. To prepare a 1 M NaCl(*aq*) solution, you could measure out 1 mol (58.5 g) of NaCl and then add water up to 1 L of solution so that the 1 mol NaCl is dissolved in the resulting 1 L solution. Note that you "add water up to 1 L" not "add up to 1 L of water," which means something entirely different. Adding water *up to* 1 L of solution won't quite require 1 L of water because the solute will displace a small amount of water. You won't always want 1 L of solution, however. To prepare a particular volume and concentration of solution requires calculating the mass of solute to weigh out. Chemists generally memorize the formula for this calculation through countless repetitions in the lab.

### Sample Problem — Converting Volume of Solution into Mass of Solute

Describe how to prepare 0.055 L of 0.20 M KCl from the solid.




#### What to Think about

1. Convert: L soln ⟶ mol KCl ⟶ g KCl

2. Setup:
$$0.055 \text{ L soln} \times \frac{? \text{ mol KCl}}{1 \text{ L soln}} \times \frac{? \text{ g KCl}}{1 \text{ mol KCl}}$$

3. Conversion factors:
0.20 mol KCl per 1 L soln
74.6 g KCl per 1 mol KCl

#### How to Do It

$$0.055 \cancel{\text{ L soln}} \times \frac{0.20 \cancel{\text{ mol KCl}}}{1 \cancel{\text{ L soln}}} \times \frac{74.6 \text{ g KCl}}{1 \cancel{\text{ mol KCl}}}$$
$= 0.82$ g KCl

Measure out 0.82 g KCl and add water up to 55 mL (0.055 L) of solution.

### Sample Problem — Determining Molar Concentration

What molar concentration of KCl is produced by measuring out 1.0 g KCl and adding water up to 0.350 L of solution?

| What to Think about | How to Do It |
|---|---|
| 1. Convert: g KCl $\rightarrow$ mol KCl | |
| 2. Setup: $1.0\text{ g KCl} \times \dfrac{1\text{ mol KCl}}{?\text{ g KCl}}$ | $1.0\ \cancel{\text{g KCl}} \times \dfrac{1\text{ mol KCl}}{74.6\ \cancel{\text{g KCl}}} = 0.013\text{ mol KCl}$ |
| 3. Conversion factor: 1 mol KCl per 74.6 g KCl | |
| 4. Molarity is moles per litre meaning the number of moles divided by the number of litres. | Molar Concentration $\text{KCl} = \dfrac{0.013\text{ mol KCl}}{0.350\text{ L soln}} = 0.038M\text{ KCl}$ |

### Practice Problems — Converting Volume of Solution into Mass of Solute and Determining Molar Concentration

1. Describe how to prepare 500.0 mL of 1.5 M $CaCl_2$ from $CaCl_2(s)$.

2. What mass of KCl would be recovered if 55 mL of 0.20 M KCl were "evaporated to dryness"?
[Hint: this is the same as asking how many grams of KCl are in 55 mL of 0.20 M KCl.]

3. What molar concentration of silver nitrate is produced by measuring out 1.8 g and then adding water to make 75 mL of solution?

## Ions in Solution

Recall from chapter 2 that ionic compounds have no net charge. The ions associate together in the ratio that results in their charges cancelling. For example:

$2Al^{3+}(aq) + 3SO_4^{2-}(aq) \rightarrow Al_2(SO_4)_3(s)$

The ionic compound is neutral because the ions have a net charge of zero: $2(3+) + 3(2-) = 0$. The ions, however, remain unchanged in the crystal. By convention, chemists don't show the charges of the ions in the formulas of ionic compounds. The charges are implicit (implied) rather than explicit (shown). When an ionic compound dissolves, the same ions that associated together to form the compound now **dissociate** (dis-associate) and travel independently through the solution. For example,

$Al_2(SO_4)_3(s) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$

### Sample Problem — Two-Step Conversion: Volume of Solution to Number of Ions

Some communities fluoridate their water to reduce tooth decay. HealthLinkBC reports that the most effective $F^-$ concentration for water supplies in B.C. is between 0.042 M and 0.053 M. How many fluoride ions would a person ingest by drinking 2.0 L of 0.047 M $F^-$?

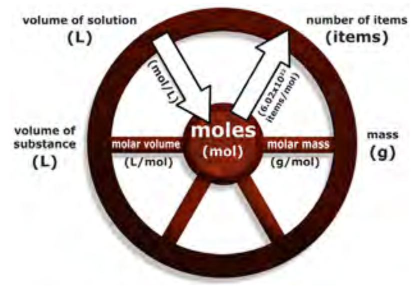
#### What to Think about

1. Convert: L soln $\rightarrow$ mol $F^-$ $\rightarrow$ ions $F^-$
2. Setup:
   $2.0\text{ L soln} \times \dfrac{?\text{ mol F}^-}{1\text{ L soln}} \times \dfrac{?\text{ ions F}^-}{1\text{ mol F}^-}$
3. Conversion factors:
   0.047 mol $F^-$ per 1 L soln
   $6.02 \times 10^{23}$ ions $F^-$ per 1 mol $F^-$

#### How to Do It

$$2.0\ \cancel{\text{L soln}} \times \frac{0.047\ \cancel{\text{mol F}^-}}{1\ \cancel{\text{L soln}}} \times \frac{6.02 \times 10^{23}\text{ ions F}^-}{1\ \cancel{\text{mol F}^-}}$$

$$= 5.7 \times 10^{22}\text{ ions F}^-$$




Being able to relate the concentration of dissolved ions to the concentration of their parent compound is extremely important in chemistry. Although it may be misleading, a label is not necessarily intended to indicate what is actually present in the solution. Some knowledge of chemistry is required to realize how a solute behaves in solution. 1 M $Al_2(SO_4)_3$ means that 1 mol of $Al_2(SO_4)_3$ was dissolved per litre of solution. There is no such thing as an $Al_2(SO_4)_3$ particle. The dissociation equation provides the ratio of the released ions to each other and to their parent compound; thus 1 M $Al_2(SO_4)_3$ actually contains 2 M $Al^{3+}$ and 3 M $SO_4^{2-}$.

### Sample Problem — Relating the Concentration of Dissolved Ions to the Concentration of Their Parent Compound

What concentrations of ions are present in 3.0 M $CaCl_2(aq)$?

#### What to Think about

The ratios in the dissociation equation show that 1 mol $Ca^{2+}$ and 2 mol $Cl^-$ are formed for each mole of $CaCl_2$ dissolved.

#### How to Do It

| **$CaCl_2(s)$** | $\rightarrow$ | **$Ca^{2+}(aq)$ +** | **$2Cl^-(aq)$** |
|---|---|---|---|
| 3.0 mol/L | dissolves to form | ? | ? |
| 3.0 mol/L | dissolves to form | 3.0 M | 6.0 M |

This table is shown for teaching purposes only — you don't need to show it in your work.

The molar concentration of a chemical is indicated by putting square brackets [ ] around the chemical's formula.

For example, $[Na^+]$ means the molar concentration of $Na^+$. A couple of precautions:

- "M" already means "mol per L" therefore *don't write* "M per L" because that would mean "moles per litre per litre," which doesn't make sense.

- You can write "2 M $Na^+$" or "$[Na^+]$ = 2 M" but *don't write* "2 M $[Na^+]$" because that would mean "two molar the molar concentration of $Na^+$," which doesn't make sense.

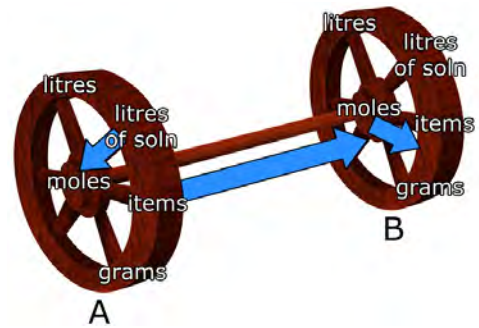
The dissociation equation provides the conversion factor represented by the axle in our wheel model.

## Sample Problem — Three-Step Conversion: Volume of Solution to Number of Ions

Aluminum chloride can be used to produce aluminum chlorohydrate, an active ingredient in antiperspirants. How many chloride ions are in 0.025 L of 0.30 M $AlCl_3$?

### What to Think about

1. Convert:
   L soln ⟶ mol $AlCl_3$ ⟶ mol $Cl^-$ ⟶ ions $Cl^-$
2. Setup: $0.025 \text{ L soln} \times \dfrac{? \text{ mol } AlCl_3}{1 \text{ L soln}} \times \dfrac{? \text{ mol } Cl^-}{1 \text{ mol } AlCl_3} \times \dfrac{? \text{ ions } Cl^-}{1 \text{ mol } Cl^-}$
3. Conversion factors:
   0.30 mol $AlCl_3$ per 1 L soln
   3 mol $Cl^-$ per 1 mol $AlCl_3$
   $6.02 \times 10^{23}$ $Cl^-$ ions per 1 mol $Cl^-$




### How to Do It

$$0.025 \cancel{\text{L soln}} \times \frac{0.30 \cancel{\text{mol } AlCl_3}}{1 \cancel{\text{L soln}}} \times \frac{3 \cancel{\text{mol } Cl^-}}{1 \text{ mol } \cancel{AlCl_3}} \times \frac{6.02 \times 10^{23} \text{ ions } Cl^-}{1 \cancel{\text{mol } Cl^-}} = 1.4 \times 10^{22} \text{ ions } Cl^-$$

## Practice Problems — Three-Step Conversion: Volume of Solution to Number of Ions

1. What concentrations of ions are present in 1.5 M $CaCl_2$(*aq*)?
2. What concentration of sodium phosphate contains 0.60 M $Na^+$?
3. Write the relationship between the concentrations of the ions present in a solution of lithium phosphate. (Careful; this is tricky. In a $CaCl_2$ solution, $[Cl^-] = 2\,[Ca^{2+}]$ ).
4. What mass of potassium ions is in 0.75 L of 2.8 M $K^+$?
5. Iron(III) nitrate solutions are used by jewellers to etch silver. How many $NO_3^-$ ions are dissolved in a 525 mL bath of 3.0 M iron(III) nitrate?

# 3.6 Activity: Building a Scale Model of a Solution

### Question
What is the ratio of solute ions to water molecules in a solution of 1 M NaCl(*aq*)?

### Background
Models are very important in science. A model, scientific or otherwise, is anything that represents something else. Chemists can use models to explain and predict the behaviour of matter. The American chemist, Linus Pauling, figured out the spiral structure of proteins using paper cut-outs as models. He was awarded the 1954 Nobel Prize in chemistry partly for this accomplishment. In this activity, you will construct a scale model of a solution.

### Procedure
1. Calculate how many moles of $H_2O$ molecules occupy one L. (Density of $H_2O$ = 1000 g/L)

$$\frac{\_\_\_\_\_\ \text{g } H_2O}{1\ \text{L } H_2O} \times \frac{1\ \text{mol } H_2O}{\text{g } H_2O} = \_\_\_\_\_\_\_\_\_\_\_\_\ \text{mol } H_2O/\text{l } H_2O$$

   Although its units are moles per litre this value is not a concentration; it is not molarity. It does not describe the proportion of a chemical in a mixture. It describes the number of moles of water in 1 L of the substance and is simply the inverse of water's molar volume.

2. Determine the ion concentrations in 1.0 M NaCl(*aq*).

| **NaCl** | → | **Na$^+$** + | **Cl$^-$** |
|---|---|---|---|
| 1.0 M | dissolves to form | | |

3. For simplicity, let's assume that each mole of ions displaces a mole of water molecules. State the ratio of water molecules: sodium ions: chloride ions in 1 M NaCl (*aq*).

   ________ $H_2O$: ________ $Na^+$: ________$Cl^-$

4. As a class, decide which kind or colour of bead will represent each chemical species.

   $H_2O$ molecules ____________________ $Na^+$ ions _______________________

   $Cl^-$ ions ______________________

5. Count out the beads in the ratio shown in step 3 and pour them into the 500-mL graduated cylinder provided by your teacher for the class.

### Results and Discussion
6. State three ways your model differs from an actual 1 M NaCl solution.

# 3.6 Review Questions

1. What does 1.5 M HCl mean?

2. A cough syrup contains 0.011 M dextromethorphan. How many moles of the cough suppressant are in one teaspoon (5.0 mL) of the cough syrup?

3. 75.0 mmol $Ca^{2+}$ = ________ L of 0.20 M $Ca^{2+}$

4. The fluid inside living cells is called cytosol. A human hepatocyte (a type of liver cell) with a volume of 500 fL (1 fL (femtolitre) = $10^{-15}$ L) contains 12 mM $Na^{+}$. How many sodium ions are in the cytosol of this cell?

5. Consumer products express concentrations in mg/volume or g/volume because the general public isn't familiar with molarity.
(a) A medium-sized (296 mL) cup of Tim Horton's coffee contains 0.10 g caffeine, $C_8H_{10}N_4O_2$. Express this concentration in molarity.

(b) A 355 mL can of pop contains 42.6 g sugar, $C_6H_{12}O_6$. What is the sugar's molar concentration?

6. Humans have an average blood volume of 5.0 L with an average blood sugar ($C_6H_{12}O_6$) concentration of 4.0 mM. What is the average mass of glucose coursing through the human bloodstream?

7. Describe how to prepare 250 mL of 0.50 M sodium nitrate. Be sure to answer in a complete sentence.

8. As a glass of cold tap water warms up, small air bubbles will come out of solution on the inner wall of the glass. A glass of cold water contains 0.45 mM $O_2$. How many millilitres of oxygen gas at STP are dissolved in 300.0 mL of this water?

9. What concentrations of ions are present in:
(a) 0.35 M $Fe_2(Cr_2O_7)_3$?

(b) 1.6 mol/L strontium cyanide?

10. In reflected light, iron(III) chloride crystals appear dark green but in transmitted light they appear maroon. What concentration of iron(III) chloride contains 0.038 M $Cl^-$?

11. In a solution of $Fe_2(SO_4)_3$:
(a) if the $[Fe^{3+}]$ = 1.5 M then what is the $[SO_4^{2-}]$?

(b) if the $[SO_4^{2-}]$ = 3.0 M then what is the $[Fe^{3+}]$?

12. Write the relationship between the concentrations of the ions in a solution of:
(a) zinc chromate

(b) strontium hydroxide

13. Milk has a $[Ca^{2+}]$ of about 31.4 mM. What mass of $Ca^{2+}$ ions are in a 250 mL serving of milk?

14. How many $Na^+$ ions are dissolved in 1.5 L of 3.0 M $Na_2CO_3$?

15. It takes 145 drops from a pipette to reach the 5.0 mL mark on a graduated cylinder. How many grams of bromide ions are in one such drop of 0.10 M iron(III) bromide?

16. Phosphoric acid, $H_3PO_4$, is added to soft drinks to increase their tartness and to act as a preservative. The concentration of $H_3PO_4$ in Pepsi is proprietary (a company secret) but can be determined from its phosphorus content since $H_3PO_4$ is the only source of phosphorus in the beverage. There are 49 mg of phosphorus in a 355 mL can of Pepsi. What is the $[H_3PO_4]$ in Pepsi?

17. Draw the plot representing a 1.5 M NaCl solution on the graph provided.

**Amount of NaCl vs. Volume of Solution**