The Mole and Avogadro's Number

Reflect

Imagine a carpenter who needs enough nails to lay 240 shingles on the roof of a new house. Because each shingle requires four nails, the carpenter multiplies 240 by 4 to find the total number of nails he needs: 960. Do you think he will count out each nail separately when he goes to the hardware store to buy supplies? Or will he buy nails already packaged in boxes?

Having the right amount of supplies on hand helps carpenters complete their work efficiently.

Which is easier and why? How do you think the method a carpenter uses to measure the number of nails is similar to the method a chemist uses to measure atoms or molecules in a chemical reaction?

The Mole: An SI Unit for Describing an Amount of a Substance

Nails are usually packaged in boxes. Each box contains a large number of nails. That's because many nails may be needed for one project, and buying them by the box is easier. Instead of worrying about individual nails, a carpenter can choose the correct number of boxes of nails needed to complete a project.

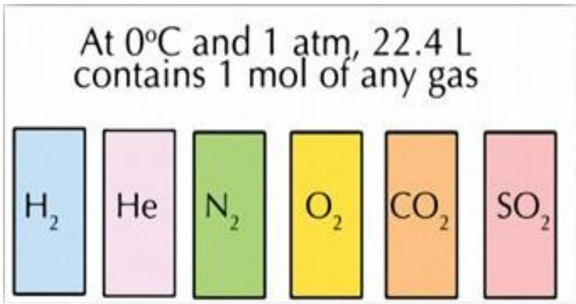
Chemists use a similar approach when they carry out chemical reactions. Instead of thinking about the individual atoms or molecules involved, they focus on groups of them. Chemists use a unit known as the *mole* (abbreviated: *mol*) to define a specific group of particles. One mole is the SI unit for measuring the quantity of matter. It is defined as the amount of a substance that contains the same number of particles as the number of atoms in exactly 12 grams of carbon-12. More simply, 1 mole equals 6.02×10^{23} particles. These particles can be atoms, molecules, ions, electrons, or any other chemical unit.

Just as a box of nails contains a specific number of nails, a mole of atoms contains a specific number of atoms.

Scientists in the Spotlight: Amedeo Avogadro

Amedeo Avogadro, an Italian scientist who lived from 1776 to 1856, developed the concept of the mole. Avogadro studied to be a lawyer and practiced law near Turin, a city in the northwestern region of Italy. However, Avogadro was fascinated with science. He engaged private tutors in mathematics and science and later gave up law to become a professor of chemistry at the University of Turin.

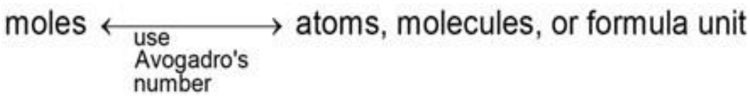
In 1811, Avogadro published a landmark paper. He had completed some experiments with different types of gases. Avogadro wrote his results in the paper along with the hypothesis that different gases contain the same number of particles whenever the conditions of volume, temperature, and pressure are the same. This was a novel idea. No one had considered comparing the actual numbers of particles in a substance before.



Avogadro’s hypothesis was mostly ignored for more than 50 years. Eventually, scientists recognized its usefulness, and Avogadro’s work became the foundation for the concept of the mole. Scientists later established that 1 mole of a gas occupies a volume of 22.4 liters when the temperature is 0 degrees Celsius and the pressure is 1 atmosphere (atm). A sample like this contains 6.02×10^{23} particles, the numerical value of 1 mole. Today, 6.02×10^{23} is known as *Avogadro’s number* in honor of Avogadro’s work.

Using Avogadro’s Number

Avogadro’s number is used as a conversion factor whenever the number of particles in a substance must be calculated.



For example, if you have 0.25 moles of hydrogen gas (H_2), you can use a conversion factor with Avogadro's number (shown in red) to determine how many molecules of hydrogen you have:

$$0.25\ \cancel{\text{mol } H_2} \times \frac{6.02 \times 10^{23}\ \text{molecules } H_2}{1\ \cancel{\text{mol } H_2}} = 1.5 \times 10^{23}\ \text{molecules } H_2$$

Alternatively, you may perform the reverse calculation if you know the number of particles in a substance and must determine the number of moles:

$$1.5 \times 10^{23}\ \cancel{\text{molecules } H_2} \times \frac{1\ \text{mol } H_2}{6.02 \times 10^{23}\ \cancel{\text{molecules } H_2}} = 0.25\ \text{mol } H_2$$

Conversions Between Moles and Mass

While the mole is a convenient measuring unit for particles, no one has invented a device to actually measure moles directly in a sample. Instead, chemists use an electronic balance to measure the mass of a sample, and then they perform a calculation to convert mass to moles. The calculation uses *molar mass* as a conversion factor:

$$\text{grams} \xleftrightarrow[\text{use molar mass}]{} \text{moles}$$

The mass of a substance can be accurately measured on an electronic balance and used to calculate the number of moles of that substance.

Molar mass is defined as the mass of 1 mole of a substance. Molar mass has units of grams/mole, and it has the same numerical value as the formula weight (or molecular weight) of the substance. Formula weight (or molecular weight) is determined by adding the atomic masses of all atoms present in one formula unit (or molecule) of the substance. An element's atomic mass—measured in atomic mass units (amu)—can be found in most copies of the periodic table.

For example, suppose you had 55.4 grams of magnesium bromide ($MgBr_2$). First, use the periodic table to determine the formula weight of $MgBr_2$.

- One formula unit of $MgBr_2$ consists of one atom of magnesium (Mg) and two atoms of bromine (Br).
- The atomic mass of Mg is 24.3 atomic mass units. The atomic mass of bromine is 79.9.
- Therefore, the formula weight of $MgBr_2$ equals 24.3 amu + (2 × 79.9 amu), or 184.1 atomic mass units.

Because a substance’s molar mass has the same numerical value as its formula weight, the molar mass of $MgBr_2$ equals 184.1 grams per mole. You can use this value as a conversion factor to determine the number of moles in a 55.4 gram sample of $MgBr_2$:

$$55.4\ \cancel{\text{g MgBr}_2} \times \frac{1\ \text{mol MgBr}_2}{184.1\ \cancel{\text{g MgBr}_2}} = 0.301\ \text{mol MgBr}$$

Alternatively, suppose you needed exactly 2.00 moles of $MgBr_2$ to carry out a chemical reaction. You can use molar mass to calculate the number of grams of $MgBr_2$ you would need to weigh to obtain the required number of moles:

$$2.00\ \cancel{\text{mol MgBr}_2} \times \frac{184.1\ \text{g MgBr}_2}{1\ \cancel{\text{mol MgBr}_2}} = 368\ \text{g MgBr}_2$$

Considering both Avogadro’s number and molar mass as conversion factors, you can convert between grams, moles, and number of particles:

$$\text{grams} \xleftrightarrow[\text{molar mass}]{\text{use}} \text{moles} \xleftrightarrow[\text{Avogadro's number}]{\text{use}} \text{atoms, molecules, or formula units}$$

What Do You Think?

How would you use molar mass and Avogadro’s number to calculate the number of atoms in 1 gram of iron?

Calculating the Percent Composition of a Compound

Molar mass is useful in other types of calculations, too. For example, chemists often want to know the elemental makeup of a compound, including the identity of all the elements present and how much of each element is present. One means for expressing this distribution is by the compound’s *percent composition,* which gives the percent by mass of each element in a compound. The general formula for calculating percent composition is shown below. Notice that molar mass is used in the denominator:

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

For example, the chemical formula of water is H_2O. This formula reveals the identities of the elements in the compound. One molecule of water consists of two hydrogen atoms (H) and one oxygen atom (O). Because water consists of two elements, the percent composition calculation will consist of two percentage values, one for each element. To find the percent composition of each element in water, first determine the mass of each element in 1 mole of water. You also need to determine the molar mass of water by summing the masses of the individual elements:

$$\text{grams of H in one mole of } H_2O\text{:} \quad 2\ \cancel{\text{mol H}} \times \frac{1.01\text{ g H}}{1\ \cancel{\text{mol H}}} = 2.02\text{ g H}$$

$$\text{grams of O in one mole of } H_2O\text{:} \quad 1\ \cancel{\text{mol O}} \times \frac{16.00\text{ g O}}{1\ \cancel{\text{mol O}}} = 16.00\text{ g O}$$

molar mass of H_2O is the sum of these two values: sum = 18.02 g/mole

Next, apply the formula for percent composition by dividing the mass of each element by the molar mass. Multiply the result by 100 to express the result as a percentage:

$$\%\text{ H in } H_2O\text{:} \quad \frac{2.02\text{ g H}}{18.02\text{ g } H_2O} \times 100 = 11.2\ \%\text{ H}$$

$$\%\text{ O in } H_2O\text{:} \quad \frac{16.00\text{ g O}}{18.02\text{ g } H_2O} \times 100 = 88.8\ \%\text{ O}$$

The sum of the percentages of all elements = 100.0%

A good way to check your work is to add the percent values. They should add up to 100%. If they don't, you've made a mistake in your calculations somewhere.

A water molecule is 88.8 percent oxygen and 11.2 percent hydrogen. (This molecule is not drawn to scale. An oxygen atom has much more mass than a hydrogen atom.)

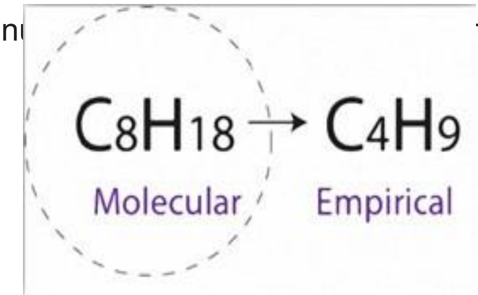
Look Out!

Though not explicitly indicated, percent composition is always based on the masses and not on the moles of elements in a compound. *Mole fraction* is used to indicate the ratio of moles of each element. It is important to avoid confusing these two types of calculations. It helps to think about the case of water as an example. In water, the mole fraction of hydrogen is very high because there are 2 moles of hydrogen for every 1 mole of oxygen. Yet because a hydrogen atom has such a small mass compared to an oxygen atom, the percent by mass of hydrogen in water is very low.

Empirical and Molecular Formulas

The chemical formula of a compound provides information about the elements present in the compound and their mole ratios. However, there are two types of chemical formulas. When chemists analyze new compounds, they often collect data on the elements making up these compounds. They then use the data to determine the following two formulas.

- **Empirical formula**: An empirical formula expresses the lowest whole number ratio of the elements in a compound.
- **Molecular formula**: A molecular formula contains the actual n[illegible]nt making up one molecule or formula unit of a compound.



Consider the compound *n*-octane. One molecule of *n*-octane has 8 carbon and 18 hydrogen atoms, giving it a molecular formula of C_8H_{18}. The ratio of carbon to hydrogen in this compound is 8:18. This is not the smallest whole number ratio because both 8 and 18 can be divided by 2 to give a ratio of 4:9.

Because the ratio 4:9 cannot be reduced further, it is the smallest whole number ratio. Therefore, the empirical formula for *n*-octane is C_4H_9.

An empirical formula expresses the simplest ratio of elements given in the molecular formula of a compound.

Often, when a new compound has been discovered, its empirical formula is determined first. This is because a chemist will begin by analyzing the identities of the elements and their percent composition. For example, suppose that an unknown compound was analyzed and found to contain 38.7% carbon (C), 9.8% hydrogen (H), and 51.5% oxygen (O). What is the empirical formula of the compound?

To solve this problem, you must assume you have a 100-gram sample of the unknown compound. This gives you 38.7 grams of carbon, 9.8 grams of hydrogen, and 51.5 grams of oxygen. Convert each of these mass measurements into moles using the molar mass of each element as a conversion factor:

$$38.7\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 3.22\ \text{mol C}$$

$$9.8\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}} = 9.70\ \text{mol H}$$

$$51.5\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 3.22\ \text{mol O}$$

Take the smallest mole quantity and divide all of the mole quantities by this number to generate mole ratios:

$$\frac{3.22 \text{ mol C}}{3.22 \text{ mol C}} = 1.00$$

$$\frac{9.70 \text{ mol H}}{3.22 \text{ mol C}} = 3.01$$

$$\frac{3.22 \text{ mol O}}{3.22 \text{ mol C}} = 1.00$$

Rounding each number to an integer, the simplest whole number ratio for C:H:O is: 1:3:1. The empirical formula for this compound is CH_3O.

Calculating a Molecular Formula

The molecular formula of a new compound is usually determined after its empirical formula has been established. In fact, a chemist can use the empirical formula of the compound, along with an experimentally determined molar mass, to find the molecular formula. For example, suppose the unknown compound above with the empirical formula CH_3O was found to have a molar mass of 62.1 grams per mole. What is the molecular formula of the compound?

You can solve this problem by comparing the experimentally determined molar mass that represents the molecular formula with a calculated molar mass based on the empirical formula. The two molar masses should be related by a whole number value.

empirical formula = CH_3O
molar mass calculated from empirical formula = 12.01 + 3(1.01) + 16.0

molecular formula = $C_?H_?O_?$
experimentally determined molar mass = 62.1 g/mol

$$\frac{\text{molar mass representing molecular formula}}{\text{molar mass representing empirical formula}} \quad \frac{62.1 \text{ g/mol}}{31.04 \text{ g/mol}} \mid 2.00$$

In this case, the two molar masses are related by a whole number factor of 2. This means the molecular formula is related to the empirical formula by a factor of 2. Because the empirical formula is CH_3O, the molecular formula must be $C_2H_6O_2$.

Try Now

What do you know?
Use your understanding of the mole and molar mass to fill in the empty cells in the following table.

Compound number	Percent composition	Molar mass (g/mol)	Empirical formula	Molecular formula	Number of molecules or formula units in 1.0 g of the compound
1	5.9% H, 94.1% O	34.02			
2		159.69	Fe_2O_3		
3				$C_{10}H_{20}C_5$	

The Mole and Avogadro's Number

Connecting With Your Child

Moles and Everyday Substances
To help your child develop a better understanding of the mole concept, pose the following challenge: A penny has a mass of 2.50 grams and is 2.5% copper (Cu). (A penny contains mostly zinc.) How many pennies are necessary to make up 1 mole of copper?

To solve this problem, you must first know the molar mass of copper. According to the periodic table, the average atomic mass of copper is approximately 63.5. Therefore, copper's molar mass is 63.5 grams per mole. How many pennies are necessary to make up 63.5 grams, or 1 mole, of copper? To solve this problem, calculate the mass of copper in one penny: 2.50 g × 0.025 = 0.0625 g. In other words, one penny contains approximately 0.0625 grams of copper, and 1,000 pennies contain approximately 62.5 grams of copper. You need approximately 1,016 pennies to make up 1 mole of copper.

Once your child has solved this problem, he or she can solve similar problems involving the composition of other coins. For example:

- A nickel is 5.00 grams and is 75% copper (Cu) and 25% nickel (Ni).
- A dime is 2.268 grams and is 91.67% copper and 8.33% nickel.
- A quarter is 5.67 grams and is 91.67% copper and 8.33% nickel.

You can apply the same principles to mixtures other than coins. For example, air is approximately 78% nitrogen and 21% oxygen. (The remaining 1% of air is made up of various other gases.) How many moles of air is necessary to make up 1 mole of oxygen?

Discuss the following questions with your child:

1. Why do scientists measure quantity of matter in moles? (What are the advantages of working in moles?)
2. Why might 1 mole of one substance have a different volume than 1 mole of a different substance?