For example, photosynthesis involves the reaction of carbon dioxide and water to make sugars. Richard Feynman noted chemistry's contribution in a very poetic way:

“The world looks so different after learning science. For example, trees are made of air, primarily. When they are burned, they go back to air, and in the flaming, heat is released, the flaming heat of the sun which was bound in to convert the air into tree. And in the ash is the small remnant of the part which did not come from air, that came from the solid earth, instead.

Sugar (and other "carbohydrates" in plants) is made photosynthetically according to overall equations like

\[\text{(6 CO}_2 + 6 \text{ H}_2\text{O} + \text{Light Energy} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{ O}_2)\]

This explains why plants die without water. But how much water is required to make a pound (457 grams) of sugar? The equation tells us 6 water molecules make 6 glucose (sugar) molecules. How do we get useful, macroscopic answers, like how much water (and CO\text{2}) is needed for a pound of sugar? We start with molar masses.

It is often convenient to express physical quantities per unit amount of substance (per mole), because in this way equal numbers of atoms or molecules are being compared. Such molar quantities often tell us something about the atoms or molecules themselves. For example, if the molar volume of one solid is larger than that of another, it is reasonable to assume that the molecules of the first substance are larger than those of the second. (Comparing the molar volumes of liquids, and especially gases, would not necessarily give the same information since the molecules would not be as tightly packed.)

A molar quantity is one which has been divided by the amount of substance. For example, an extremely useful molar quantity is the molar mass \(M\):

\[\text{Molar mass} = \frac{\text{mass}}{\text{amount of substance}}\]

**EXAMPLE 1** Obtain the molar mass of (a) C (carbon) and (b) H\text{2O}.

**Solution**

a) The atomic weight of carbon is 12.01, and so 1 mol C weighs 12.01 g.
\[ M_{\text{C}} = \frac{m_{\text{C}}}{n_{\text{C}}} = \frac{12.01 \text{ g}}{1 \text{ mol}} = 12.01 \text{ g mol}^{-1} \]

b)

\[ M_{\text{H}_2\text{O}} = \frac{m_{\text{H}_2\text{O}}}{n_{\text{H}_2\text{O}}} = 18.02 \text{ g mol}^{-1} \]

density Avogadro constant via

\[ \text{Mass} \overset{\text{Molar mass}}{\longleftrightarrow} \text{amount of substance} \]

\[ m \overset{M}{\longleftrightarrow} n \]

The molar mass is easily obtained from atomic weights and may be used as a conversion factor, provided the units cancel.

**EXAMPLE 2** Calculate the amount of glucose (C\(_6\)H\(_{12}\)O\(_6\)) in 457 g of this solid.

**Solution** Any problem involving interconversion of mass and amount of substance requires molar mass

\[ M = (6 \times 12.01 + 12 \times 1.008 + 6 \times 16) \text{ g mol}^{-1} = 180 \text{ g mol}^{-1} \]

\[ n = m \cdot \text{conversion factor} = m \cdot \frac{1}{M} = 457 \text{ g} \cdot \frac{1 \text{ mol}}{180 \text{ g}} = 2.54 \text{ mol} \]

**EXAMPLE 3** The chemical equation above tells us that 6 glucose molecules require 6 water molecules, so 6 moles of glucose require 6 moles of water, and since they’re equal, the 2.54 moles of glucose above would require 2.54 moles of water. What mass of water is that?

The mass of substance will be the amount times a conversion factor which permits cancellation of units:

\[ m = n \cdot \text{conversion factor} = n \cdot M = 2.54 \text{ mol} \cdot 18.02 \frac{\text{g}}{\text{mol}} = 45.7 \text{ g} \]

**EXAMPLE 4** How many molecules would be present in 50 mL of pure water?

**Solution** In previous examples, we showed that the number of molecules may be obtained from the amount of substance by using the Avogadro constant. The amount of substance may be obtained from mass by using the molar mass, and mass from volume by means of density. A road map to the solution of this problem is

\[ \text{Volume} \xrightarrow{\text{density}} \text{mass} \overset{\text{Molar mass}}{\longleftrightarrow} \text{amount} \overset{\text{Avogadro constant}}{\longleftrightarrow} \text{number of molecules} \]

\[ V \xrightarrow{\rho} m \overset{M}{\longleftrightarrow} n \xrightarrow{N_A} N \]

\[ \rho = 1.0 \text{ g cm}^{-3} \]
Table of Atomic Weights

\[ M = (2 \times 1.008 + 1 \times 16.00) \text{ g mol}^{-1} = 18.02 \text{ g mol}^{-1} \]

\[ N_A = 6.022 \times 10^{23} \text{ mol}^{-1} \]

\[
\begin{align}
N &= \text{50.0 cm}^3 \cdot \frac{1.00 \text{ g}}{1 \text{ cm}^3} \cdot \frac{1 \text{ mol}}{18.02 \text{ g}} \cdot \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \\
&= 1.67 \times 10^{24} \text{ molecules}
\end{align}
\]

Notice that in this problem we had to combine techniques from previous examples. To do this you must remember relationships among quantities. For example, a volume was given, and we knew it could be converted to the corresponding mass by means of density, and so we looked up the density in a table. By writing a road map, or at least seeing it in your mind’s eye, you can keep track of such relationships, determine what conversion factors are needed, and then use them to solve the problem.

**Example 5** A student’s body may contain \(10^{27}\) water molecules and would need to accumulate about \(10^{18}\) (ten million million) water molecules per second over 18 years.

a. Show that the number of water molecules is about right, assuming a body weight of 150 lb and that water is about 70% of body weight.

b. Show that the rate of accumulation of water molecules per second is approximately right.

**Solution**

a. \[
150 \text{ lb} \cdot \frac{0.453 \text{ kg}}{1 \text{ lb}} \cdot 0.70 = 47.6 \text{ kg water}
\]

\[
\frac{47,600 \text{ g water}}{18 \text{ g mol}} = 2.6 \times 10^3 \text{ mol}
\]

\[
\frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.6 \times 10^{27} \text{ molecules}
\]

b. \[
10^{27} \text{ water molecules} \cdot \frac{18 \text{ g}}{365 \text{ d}} \cdot \frac{24 \text{ h}}{1 \text{ d}} \cdot \frac{60 \text{ min}}{1 \text{ h}} \cdot \frac{60 \text{ sec}}{1 \text{ min}} = 4 \times 10^{18} \text{ molecules per second}
\]

**Contributors**

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.