2.11: The Molar Mass

As we saw in The Amount of Substance: Moles, there is no relationship between the mass or volume of a substance and the number of molecules. But we then defined *the amount of substance*, n, to represent the number of particles. The *amount* is therefore useful in determining how much of each substance will react. While 1 g of Hg, or 1 cm^3 of Hg, reacts with a mass or volume of Br_2 that is not related to the coefficients of the chemical equation, 1 mol of Hg always reacts with 1 mol of Br_2, since one atom of Hg reacts with one molecule of Br_2:

$1\ Hg(l)$	$+1 Br_2(l)$	$1 HgBr_2(s)$
1 atom	1 molecule	1 molecule
1.00 g	0.797 g	1.797 g
1.00 cm^3	3.47 cm^3	0.30 cm^3
1.00 mol Hg $6.022 \cdot 10^{23}$	1.00 mol Br_2 $6.022 \cdot 10^{23}$	1.00 mol $HgBr_2$ $6.022 \cdot 10^{23}$

What we need is a convenient way to convert *masses* to *amounts*, and the necessary conversion factor is called the *molar mass*. 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*:

$$\text{Molar mass} = \frac{\text{mass}}{\text{amount of substance}}$$

$$M(\text{g/mol}) = \frac{m\,(\text{g})}{n\,(\text{mol})}$$

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.)

It is almost trivial to obtain the molar mass, since atomic and molecular weights expressed in grams give us the masses of 1 mol of substance.

Example 2.11.1: Molar Mass

Obtain the molar mass of (a) Hg and (b) Hg_2Br_2.

Solution

a) The atomic weight of mercury is 200.59, and so 1 mol Hg weighs 200.59 g.

$$M_{\text{Hg}} = \frac{m_{\text{Hg}}}{n_{\text{Hg}}} = \frac{200.59\text{ g}}{1\text{ mol}} = 200.59\text{ g mol}^{-1}$$

b) Similarly, for Hg_2Br_2 the molecular weight is 560.98, and so

$$\mathrm{M_{Hg_2Br_2}} = \frac{\mathrm{m_{Hg_2Br_2}}}{\mathrm{n_{Hg_2Br_2}}} = 560.98\text{ g mol}^{-1}$$

The molar mass is numerically the same as the atomic or molecular weight, but it has units of grams per mole. The equation, which defines the molar mass, has the same form as those defining density, and the Avogadro constant. As in the case of density or the Avogadro constant, it is not necessary to memorize or manipulate a formula. Simply remember that mass and amount of substance are related *via* molar mass.

$$\text{Mass} \overset{\text{Molar mass}}{\longleftrightarrow} \text{amount of substance } m \overset{M}{\longleftrightarrow} n$$

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

Example 2.11.2: Moles

Calculate the amount of octane (C_8H_{18}) in 500 g of this liquid.

Solution

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

$$M = 8 \cdot 12.01 + 18 \cdot 1.008 \text{ g mol}^{-1} = 114.2 \text{ g mol}^{-1}$$

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

$$n = m \cdot \text{conversion factor} = \text{m} \cdot \frac{1}{M} = 500 \text{ g} \cdot \frac{1 \text{ mol}}{114.2 \text{ g}} = 4.38 \text{ mol}$$

In this case the reciprocal of the molar mass was the appropriate conversion factor.

The Avogadro constant, molar mass, and density may be used in combination to solve more complicated problems.

Example 2.11.3: Molecules

How many molecules would be present in 25.0 ml of pure carbon tetrachloride (CCl_4)?

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:

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

or in shorthand notation:

$$\text{V} \xrightarrow{\rho} m \xrightarrow{M} n \xrightarrow{N_A} N$$

The road map tells us that we must look up the density of CCl_4(opens in new window):

$$\rho = 1.595 \text{ g cm}^{-3}$$

The molar mass must be calculated from the Table of Atomic Weights(opens in new window).

$$M = (12.01 + 4 \cdot 35.45) \text{ g mol}^{-1} = 153.81 \text{ g mol}^{-1}$$

and we recall that the Avogadro constant is:

$$N_A = 6.022 \cdot 10^{23} \text{ mol}^{-1}$$

The last quantity (N) in the road map can then be obtained by starting with the first (V) and applying successive conversion factors:

$$N = 25 \text{ cm}^3 \cdot \frac{1.595 \text{ g}}{1 \text{ cm}^3} \cdot \frac{1 \text{ mol}}{153.81 \text{ g}} \cdot \frac{6.022 \cdot 10^{23} \text{ molecules}}{1 \text{ mol}} \tag{2.11.1}$$

$$= 1.56 \cdot 10^{23} \text{ molecules} \tag{2.11.2}$$

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.

This page titled 2.11: The Molar Mass is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Ed Vitz, John W. Moore, Justin Shorb, Xavier Prat-Resina, Tim Wendorff, & Adam Hahn.