1.2 THE MOLE UNIT

Your objectives in studying this section are to be able to:

1. Define a kilogram mole, pound mole, and gram mole.
2. Convert from moles to mass and the reverse for any chemical compound given the molecular weight.
3. Calculate molecular weights from the molecular formula.

What is a mole? The best answer is that a **mole** is a certain number of molecules, atoms, electrons, or other specified types of particles.[4] In particular, the 1969 International Committee on Weights and Measures approved the mole (symbol mol in the SI system) as being "the amount of a substance that contains as many elementary entities as there are atoms in 0.012 kg of carbon 12." Thus in the SI system the mole contains a different number of molecules than it does in the American engineering system. In the SI system a mole has about 6.023×10^{23} molecules; we shall call this a *gram mole* (symbol *g mol*) to avoid confusion even though in the SI system of units the official designation is simply mole (abbreviated mol). We can thereby hope to avoid the confusion that could occur with the American engineering system *pound mole* (abbreviated *lb mol*), which has $6.023 \times 10^{23} \times 453.6$ molecules. Thus a pound mole of a substance has more mass than does a gram mole of the substance.

Here is another way to look at the mole unit. To convert the number of moles to mass, we make use of the **molecular weight**—the mass per mole:

$$\text{the g mol} = \frac{\text{mass in g}}{\text{molecular weight}} \tag{1.4}$$

$$\text{the lb mol} = \frac{\text{mass in lb}}{\text{molecular weight}} \tag{1.5}$$

or

$$\text{mass in g} = (\text{mol. wt.})(\text{g mol}) \tag{1.6}$$

$$\text{mass in lb} = (\text{mol. wt.})(\text{lb mol}) \tag{1.7}$$

Furthermore, there is no reason why you cannot carry out computations in terms of ton moles, kilogram moles, or any corresponding units if they are defined analogously to Eqs. (1.4) and (1.5) even if they are not standard units. If you read

[4] For a discussion of the mole concept, refer to the series of articles in *J. Chem. Educ.*, v. 38, pp. 549–556 (1961), and to M. L. McGlashan, *Phys. Educ.*, p. 276 (July 1977).

about a unit such as a kilomole (kmol) without an associated mass specification, or a kg mol, assume that it refers to the SI system and 10^3 g mol.

Values of the molecular weights (relative molar masses) are built up from the tables of atomic weights based on an arbitrary scale of the relative masses of the elements. **Atomic weight** of an element is the mass of an atom based on the scale that assigns a mass of exactly 12 to the carbon isotope ^{12}C, whose nucleus contains 6 protons and 6 neutrons. The terms atomic "weight" and molecular "weight" are universally used by chemists and engineers instead of the more accurate terms atomic "mass" or molecular "mass." Since weighing was the original method for determining the comparative atomic masses, as long as they were calculated in a common gravitational field, the relative values obtained for the atomic "weights" were identical with those of the atomic "masses."

Appendix B lists the atomic weights of the elements. On this scale of atomic weights, hydrogen is 1.008, carbon is 12.01, and so on. (In most of our calculations we shall round these off to 1 and 12, respectively, for convenience.)

A compound is composed of more than one atom, and the molecular weight of the compound is nothing more than the sum of the weights of the atoms of which it is composed. Thus H_2O consists of 2 hydrogen atoms and 1 oxygen atom, and the molecular weight of water is $(2)(1.008) + 16.000 = 18.02$. These weights are all relative to the ^{12}C atom as 12.0000, and you can attach any unit of mass you desire to these weights; for example, H_2 can be 2.016 g/g mol, 2.016 lb/lb mol, 2.016 ton/ton mol, and so on. Gold (Au) is 196.97 oz/oz mol!

You can compute average molecular weights for mixtures of constant composition even though they are not chemically bonded if their compositions are known accurately. Later (Example 1.13) we show how to calculate the average molecular weight of air. Of course, for a material such as fuel oil or coal whose composition may not be exactly known, you cannot determine an exact molecular weight, although you might estimate an approximate average molecular weight good enough for engineering calculations. Keep in mind that the symbol lb refers to lb_m unless otherwise stated.

EXAMPLE 1.9 Use of Molecular Weights

If a bucket holds 2.00 lb of NaOH (mol. wt. = 40.0), how many

(a) Pound moles of NaOH does it contain?
(b) Gram moles of NaOH does it contain?

Solution

Basis: 2.00 lb of NaOH

(a)
$$\frac{2.00 \text{ lb NaOH}}{} \left| \frac{1 \text{ lb mol NaOH}}{40.0 \text{ lb NaOH}} \right. = 0.050 \text{ lb mol NaOH}$$

(b_1)
$$\frac{2.00 \text{ lb NaOH}}{} \left| \frac{1 \text{ lb mol NaOH}}{40.0 \text{ lb NaOH}} \right| \frac{454 \text{ g mol}}{1 \text{ lb mol}} = 22.7 \text{ g mol}$$

or

$$(\mathbf{b_2})\qquad \frac{2.00 \text{ lb NaOH}}{} \left| \frac{454 \text{ g}}{1 \text{ lb}} \right| \frac{1 \text{ g mol NaOH}}{40.0 \text{ g NaOH}} = 22.7 \text{ g mol}$$

EXAMPLE 1.10 Use of Molecular Weights

How many pounds of NaOH are in 7.50 g mol of NaOH?

Solution

Basis: 7.50 g mol of NaOH

$$\frac{7.50 \text{ g mol NaOH}}{} \left| \frac{1 \text{ lb mol}}{454 \text{ g mol}} \right| \frac{40.0 \text{ lb NaOH}}{1 \text{ lb mol NaOH}} = 0.661 \text{ lb NaOH}$$

Self-Assessment Test

1. What is the molecular weight of acetic acid (CH_3COOH)?
2. What is the difference between a kilogram mole and a pound mole?
3. Convert 39.8 kg of NaCl per 100 kg of water to kilogram moles of NaCl per kilogram mole of water.
4. How many pound moles of $NaNO_3$ are there in 100 lb?
5. One pound mole of CH_4 per minute is fed to a heat exchanger. How many kilograms is this per second?

Thought Problem

1. There is twice as much copper in 480 g of copper as there is in 240 g of copper, but is there twice as much copper in 480 g of copper as there is silver in 240 g of silver?

1.3 CONVENTIONS IN METHODS OF ANALYSIS AND MEASUREMENT

Your objectives in studying this section are to be able to:

1. Define density and specific gravity.
2. Calculate the density of a liquid or solid given its specific gravity and the reverse.
3. Look up the density or specific gravity of a liquid or solid in reference tables.
4. Interpret the meaning of specific gravity data taken from reference tables.
5. Specify the common reference material(s) used to determine the specific gravity of liquids and solids.

6. Convert the composition of a mixture from mole fraction (or percent) to mass (weight) fraction (or percent) and the reverse.
7. Transform a material from one measure of concentration to another, including mass/volume, moles/volume, ppm, and molarity.
8. Calculate the mass or number of moles of each component in a mixture given the percent (or fraction) composition, and the reverse, and compute the average molecular weight.
9. Convert a composition given in mass (weight) percent to mole percent, and the reverse.

There are certain definitions and conventions which we mention at this point since they will be used constantly throughout the book. If you memorize them now, you will immediately have a clearer perspective and save considerable trouble later.

1.3-1 Density

Density is the ratio of mass per unit volume, as, for example, kg/m^3 or lb/ft^3. It has both a numerical value and units. To determine the density of a substance, you must find both its volume and its mass. If the substance is a solid, a common method to determine its volume is to displace a measured quantity of inert liquid. For example, a known weight of a material can be placed into a container of liquid of known weight and volume, and the final weight and volume of the combination measured. The density (or specific gravity) of a liquid is commonly measured with a hydrometer (a known weight and volume is dropped into the liquid and the depth to which it penetrates into the liquid is noted) or a Westphal balance (the weight of a known slug is compared in the unknown liquid with that in water). Gas densities are quite difficult to measure; one device used is the Edwards balance, which compares the weight of a bulb filled with air to the same bulb when filled with the unknown gas.

In most of your work using liquids and solids, density will not change very much with pressure, but for precise measurements for common substances you can always look up in a handbook the variation of density with pressure. The change in density with temperature is illustrated in Fig. 1.1 for liquid water and liquid ammo-

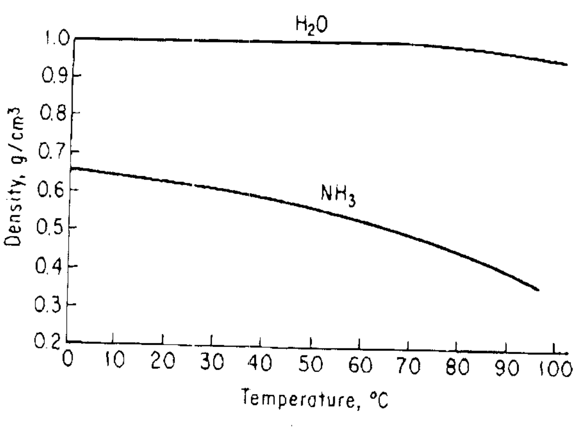

Figure 1.1 Densities of liquid H_2O and NH_3 as a function of temperature.

CHAPTER 1

Problem 1.1A

Convert the following quantities to the ones designated :

a. 42 ft^2/hr to cm^2/s.
b. 25 psig to psia.
c. 100 Btu to hp-hr.

Solution

a. $$\frac{42.0\ ft^2}{hr}\left(\frac{1.0\ m}{3.2808\ ft}\right)^2\frac{10^4 cm^2}{1.0\ m^2}\ \frac{1\ hr}{3600\ s} = \mathbf{10.8\ cm^2/s}$$

b. $$100\ Btu\ \frac{3.93\times 10^{-4}\ hp\text{-}hr}{1\ Btu} = \mathbf{3.93 \times 10^{-2}\ hp\text{-}hr}$$

c. $$\frac{80.0\ lb_f}{}\ \frac{32.174\ (lb_m)(ft)}{(lb_f)(s)^2}\ \frac{1\ kg}{2.20\ lb_m}\ \frac{1\ m}{3.2808\ ft}\ \frac{1\ N}{1\ (kg)(m)(s)^{-2}} = \mathbf{356\ N}$$

Problem 1.1 B

Convert the ideal gas constant : $R = 1.987\ \frac{cal}{(gmol)(K)}$ to $\frac{Btu}{(lb\ mol)(°R)}$

Solution

$$\frac{1.987\ cal}{(gmol)(K)}\ \frac{1\ Btu}{252\ cal}\ \frac{454\ gmol}{1\ lb\ mol}\ \frac{1K}{1.8\ °R} = \mathbf{1.98\ \frac{Btu}{(lb\ mol)(°R)}}$$

Problem 1.1 C

Mass flow through a sonic nozzle is a function of gas pressure and temperature. For a given pressure p and temperature T, mass flow rate through the nozzle is given by

$m = 0.0549\ p/(T)^{0.5}$ where m is in lb/min, p is in psia and T is in °R

a. Determine what the units for the constant 0.0549 are.
b. What will be the new value of the constant, now given as 0.0549, if the variables in the equation are to be substituted with SI units and m is calculated in SI units.

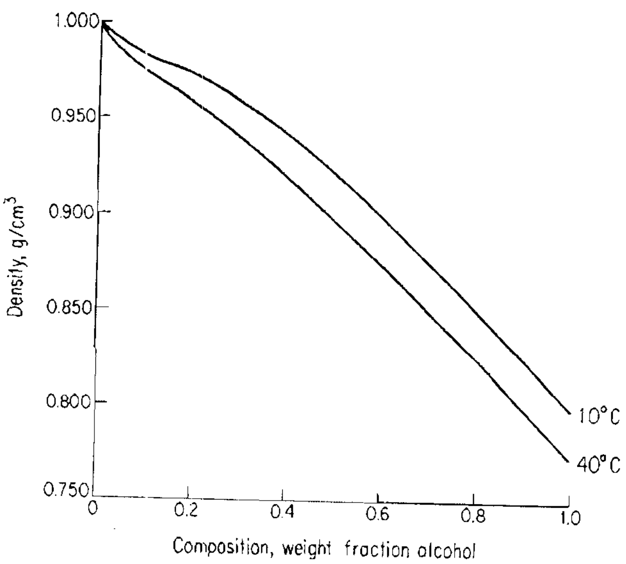

Figure 1.2 Density of a mixture of ethyl alcohol and water as a function of composition.

nia. Figure 1.2 illustrates how density also varies with composition. In the winter you may put antifreeze in your car radiator. The service station attendant checks the concentration of antifreeze by measuring the specific gravity and, in effect, the density of the radiator solution after it is mixed thoroughly. He has a little thermometer in his hydrometer kit in order to be able to read the density corrected for temperature.

1.3-2 Specific Gravity

Specific gravity is commonly thought of as a dimensionless ratio. Actually, it should be considered as the ratio of two densities—that of the substance of interest, A, to that of a reference substance. In symbols:

$$\text{sp gr} = \textbf{specific gravity} = \frac{(\text{lb/ft}^3)_A}{(\text{lb/ft}^3)_{\text{ref}}} = \frac{(\text{g/cm}^3)_A}{(\text{g/cm}^3)_{\text{ref}}} = \frac{(\text{kg/m}^3)_A}{(\text{kg/m}^3)_{\text{ref}}} \tag{1.8}$$

The reference substance for liquids and solids is normally water. Thus the specific gravity is the ratio of the density of the substance in question to the density of water. The specific gravity of gases frequently is referred to air, but may be referred to other gases, as discussed in more detail in Chap. 3. Liquid density can be considered to be nearly independent of pressure for most common calculations, but, as indicated in Fig. 1.1 it varies somewhat with temperature; therefore, to be very precise when referring to specific gravity, state the temperature at which each density is chosen. Thus

$$\text{sp gr} = 0.73\frac{20^\circ}{4^\circ}$$

can be interpreted as follows: the specific gravity when the solution is at 20°C and the reference substance (water) is at 4°C is 0.73. Since the density of water at 4°C is very close to 1.0000 g/cm^3 in the SI system, the numerical values of the specific gravity and density in this system are essentially equal. Since densities in the American engineering system are expressed in lb/ft^3 and the density of water is about 62.4 lb/ft^3, it can be seen that the specific gravity and density values are not numerically equal in the American engineering system.

In the petroleum industry the specific gravity of petroleum products is usually reported in terms of a hydrometer scale called °API. The equation for the API scale is

$$°\text{API} = \frac{141.5}{\text{sp gr}\,\frac{60°}{60°}} - 131.5 \tag{1.9}$$

or

$$\text{sp gr}\,\frac{60°}{60°} = \frac{141.5}{°\text{API} + 131.5} \tag{1.10}$$

The volume and therefore the density of petroleum products vary with temperature, and the petroleum industry has established 60°F as the standard temperature for volume and API gravity. The °API is being phased out as SI units are accepted for densities.

There are many other systems of measuring density and specific gravity that are somewhat specialized; for example, the Baumé (°Be) and the Twaddell (°Tw) systems. Relationships among the various systems of density may be found in standard reference books.

1.3-3 Specific Volume

The specific volume of any compound is the inverse of the density, that is, the volume per unit mass or unit amount of material. Units of specific volume might be ft^3/lb$_m$, ft^3/lb mole, cm^3/g, bbl/lb$_m$, m^3/kg, or similar ratios.

1.3-4 Mole Fraction and Mass (Weight) Fraction

Mole fraction is simply the moles of a particular substance divided by the total number of moles present. This definition holds for gases, liquids, and solids. Similarly, the **mass (weight) fraction** is nothing more than the **mass (weight)** of the substance divided by the total mass (weight) of all substances present. Although the mass fraction is what is intended to be expressed, ordinary engineering usage employs the term **weight fraction.** Mathematically, these ideas can be expressed as

$$\text{mole fraction of } A = \frac{\text{moles of } A}{\text{total moles}} \tag{1.11}$$

$$\text{mass (weight) fraction of } A = \frac{\text{mass (weight) of } A}{\text{total mass (weight)}} \tag{1.12}$$

Mole percent and weight percent are the respective fractions times 100.

1.3-5 Analyses

The analyses of gases such as air, combustion products, and the like **are usually on a dry basis**—that is, water vapor is excluded from the analysis. Such an analysis, called an **Orsat analysis,** is explained in Sec. 2.3. If the gas acts as an ideal gas and its components are each measured by volume, you know (or will learn in Sec. 3.1) that volume percent under ordinary conditions is the same as mole percent. For example, consider the composition of air, which is approximately

21%	oxygen
79%	nitrogen
100%	total

This means that at room temperature and pressure, any sample of air will contain 21% oxygen by volume and also be 21 mole % oxygen.

In this book, the composition of gases will always be presumed to be given in mole percent or fraction unless otherwise stated.

Analysis of liquids and solids are usually given by mass (weight) percent or fraction, but occasionally by mole percent.

In this text, the analysis of liquids and solids will always be assumed to be weight percent unless otherwise stated.

EXAMPLE 1.11 Mole Fraction and Mass (Weight) Fraction

An industrial-strength drain cleaner contains 5.00 kg of water and 5.00 kg of NaOH. What are the mass (weight) fraction and mole fraction of each component in the bottle of water?

Solution

Basis: 10.0 kg of total solution

Component	kg	Weight fraction	Mol. wt.	kg mol	Mole fraction
H_2O	5.00	$\frac{5.00}{10.0} = 0.500$	18.0	0.278	$\frac{0.278}{0.403} = 0.690$
NaOH	5.00	$\frac{5.00}{10.0} = 0.500$	40.0	0.125	$\frac{0.125}{0.403} = 0.310$
Tota.	10.00	1.000		0.403	1.00

The kilogram moles are calculated as follows:

$$\frac{5.00 \text{ kg } H_2O}{} \left| \frac{1 \text{ kg mol } H_2O}{18.0 \text{ kg } H_2O} \right. = 0.278 \text{ kg mol } H_2O$$

$$\frac{5.00 \text{ kg NaOH}}{} \left| \frac{1 \text{ kg mol NaOH}}{40.0 \text{ kg NaOH}} \right. = 0.125 \text{ kg mol NaOH}$$

Adding these quantities together gives the total kilogram moles.

EXAMPLE 1.12 Density and Specific Gravity

If dibromopentane (DBP) has a specific gravity of 1.57, what is its density in (a) g/cm^3? (b) lb_m/ft^3? and (c) kg/m^3?

Solution

No temperatures are cited for the dibromopentane or the reference compound (presumed to be water); hence we assume that the temperatures are the same and that water has a density of 1.00×10^3 kg/m^3 (1.00 g/cm^3).

(a)
$$\frac{1.57 \dfrac{\text{g DBP}}{\text{cm}^3}}{1.00 \dfrac{\text{g H}_2\text{O}}{\text{cm}^3}} \left| 1.00 \frac{\text{g H}_2\text{O}}{\text{cm}^3} \right. = 1.57 \frac{\text{g DBP}}{\text{cm}^3}$$

(b)
$$\frac{1.57 \dfrac{\text{lb}_m \text{ DBP}}{\text{ft}^3}}{1.00 \dfrac{\text{lb}_m \text{ H}_2\text{O}}{\text{ft}^3}} \left| 62.4 \frac{\text{lb}_m \text{ H}_2\text{O}}{\text{ft}^3} \right. = 97.97 \frac{\text{lb}_m \text{ DBP}}{\text{ft}^3}$$

(c)
$$\frac{1.57 \text{ g DBP}}{\text{cm}^3} \left| \left(\frac{100 \text{ cm}}{1 \text{ m}}\right)^3 \right| \frac{1 \text{ kg}}{1000 \text{ g}} = 1.57 \times 10^3 \frac{\text{kg DBP}}{\text{m}^3}$$

or

$$\frac{\dfrac{1.57 \text{ kg DBP}}{\text{m}^3}}{\dfrac{1.00 \text{ kg H}_2\text{O}}{\text{m}^3}} \left| \frac{1.00 \times 10^3 \text{ kg H}_2\text{O}}{\text{m}^3} \right. = 1.57 \times 10^3 \frac{\text{kg DBP}}{\text{m}^3}$$

Note how the units of specific gravity as used here clarify the calculations.

EXAMPLE 1.13 Average Molecular Weight of Air

What is the average molecular weight of air and its composition by weight percent?

Solution

Basis: 100 lb mol of air

Component	Moles = percent	Mol. wt.	lb	weight %
O_2	21.0	32	672	23.17
N_2*	79.0	28.2	2228	76.83
Total	100		2900	100.00

*Includes Ar, CO_2, Kr, Ne, and Xe, and is called atmospheric nitrogen. The molecular weight is 28.2.

Table 1.5 lists the detailed composition of air.

The average molecular weight is 2900 lb/100 lb mol = 29.00.

TABLE 1.5 Composition of Clean, Dry Air near Sea Level

Component	Percent by volume = mole percent
Nitrogen	78.084
Oxygen	20.9476
Argon	0.934
Carbon dioxide	0.0314
Neon	0.001818
Helium	0.000524
Methane	0.0002
Krypton	0.000114
Nitrous oxide	0.00005
Hydrogen	0.00005
Xenon	0.0000087
Ozone	
Summer	0–0.000007
Winter	0–0.000002
Ammonia	0–trace
Carbon monoxide	0–trace
Iodine	0–0.000001
Nitrogen dioxide	0–0.000002
Sulfur dioxide	0–0.0001

Do **not** attempt to get an average specific gravity or average density for a mixture of solids or liquids by multiplying the individual component specific gravities or densities by the respective mole fractions of the components in the mixture and summing the products. The proper way to use specific gravity is demonstrated in the next example.

EXAMPLE 1.14 Application of Specific Gravity

In the production of a drug having a molecular weight of 192, the exit stream from the reactor flows at the rate of 10.3 L/min. The drug concentration is 41.2% (in water), and the specific gravity of the solution is 1.025. Calculate the concentration of the drug (in kg/L) in the exit stream, and the flow rate of the drug in kg mol/min.

Solution

For the first part of the problem, we want to transform the mass fraction of 0.412 into mass per liter of the drug. Take 1.000 kg of exit solution as a basis for convenience. See Fig. E1.14.

Basis: 1.000 kg solution

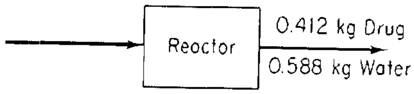

Figure E1.14

How do we get mass per volume (the density) from the given data which is in terms of mass per mass? Use the specific gravity of the solution.

$$\text{density of solution} = \frac{1.025 \dfrac{\text{g soln}}{\text{cm}^3}}{1.000 \dfrac{\text{g } H_2O}{\text{cm}^3}} \left| \; 1.000 \frac{\text{g } H_2O}{\text{cm}^3} \right. = 1.025 \frac{\text{g soln}}{\text{cm}^3}$$

Next

$$\frac{0.412 \text{ kg drug}}{1.000 \text{ kg soln}} \left| \frac{1.025 \text{ g soln}}{1 \text{ cm}^3} \right| \frac{1 \text{ kg}}{10^3 \text{ g}} \left| \frac{10^3 \text{ cm}^3}{1 \text{ L}} \right. = 0.422 \text{ kg drug/L soln}$$

To get the flow rate, we take a different basis, 1 minute.

Basis: 1 min $\equiv$ 10.3 L solution

$$\frac{10.3 \text{ L soln}}{1 \text{ min}} \left| \frac{0.422 \text{ kg drug}}{\text{L soln}} \right| \frac{1 \text{ kg mol drug}}{192 \text{ kg drug}} = 0.0226 \text{ kg mol/min}$$

1.3-6 Concentrations

Concentration means the quantity of some solute per fixed amount of solvent, or solution, in a mixture of two or more components; for example:

(a) Mass per unit volume (lb_m of solute/ft^3, g of solute/L, lb_m of solute/bbl, kg of solute/m^3).

(b) Moles per unit volume (lb mol of solute/ft^3, g mol of solute/liter, g mol of solute/cm^3).

(c) Parts per million—a method of expressing the concentration of extremely dilute solutions. Ppm is equivalent to a weight fraction for solids and liquids because the total amount of material is of a much higher order of magnitude than the amount of solute; it is a mole fraction for gases. Why?

(d) Other methods of expressing concentration with which you should be familiar are molarity (g mole/liter) and normality (equivalents/liter).

A typical example of the use of some of these concentration measures is the set of guidelines by which the Environmental Protection Agency defined the extreme levels at which the five most common air pollutants could harm people over stated periods of time.

(a) *Sulfur dioxide:* 365 $\mu g/m^3$ averaged over a 24-hr period

(b) *Particulate matter:* 260 $\mu g/m^3$ averaged over a 24-hr period

(c) *Carbon monoxide:* 10 mg/m^3 (9 ppm) when averaged over an 8-hr period; 40 mg/m^3 (35 ppm) when averaged over 1 hr

(d) *Nitrogen dioxide:* 100 μg/m^3 averaged over 1 year

It is important to remember that in an ideal solution, such as in gases or in a simple mixture of hydrocarbon liquids or compounds of like chemical nature, the volumes of the components may be added without great error to get the total volume of the mixture. For the so-called nonideal mixtures this rule does not hold, and the total volume of the mixture is bigger or smaller than the sum of the volumes of the pure components.

In Chap. 2 we will use **stream flows** and compositions in making material balances. To calculate the **mass flow rate,** $\dot{m}$, from a known **volumetric flow rate,** q, you multiply the volumetric flow rate by the mass concentration thus

$$\frac{q\ \text{m}^3}{\text{s}}\left|\frac{\rho\ \text{kg}}{\text{m}^3}\right. = \dot{m}\frac{\text{kg}}{\text{s}}$$

How would you calculate the volumetric flow rate from a known mass flow rate? From the volumetric flow rate you can calculate the **average velocity,** v, in a pipe if you know the area, A, of the pipe from the relation

$$q = Av$$

Self-Assessment Test

1. Answer the following questions true (T) or false (F).
   (a) The density and specific gravity of mercury are the same.
   (b) The inverse of density is the specific volume.
   (c) Parts per million expresses a mole ratio.
   (d) Concentration of a component in a mixture does not depend on the amount of the mixture.
2. A cubic centimeter of mercury has a mass of 13.6 g at the earth's surface. What is the density of the mercury?
3. What is the approximate density of water at room temperature?
4. For liquid HCN, a handbook gives: sp gr 10°C/4°C = 1.2675. What does this mean?
5. For ethanol, a handbook gives: sp gr 60°F/60°F = 0.79389. What is the density of ethanol at 60°F?
6. Commercial sufuric acid is 98% H_2SO_4 and 2% H_2O. What is the mole ratio of H_2SO_4 to H_2O?
7. A container holds 1.704 lb of HNO_3/lb of H_2O and has a specific gravity of 1.382 at 20°C. Compute the composition in the following ways:
   (a) Weight percent HNO_3
   (b) Pounds HNO_3 per cubic foot of solution at 20°C
   (c) Molarity at 20°C

a. $\dfrac{12.7 \text{ kg sugar}}{100 \text{ kg soln}} \left| \dfrac{100 \text{ kg solution}}{87.3 \text{ kg } H_2O} \right. = .145 \dfrac{\textbf{kg sugar}}{\textbf{kg } \mathbf{H_2O}}$

b. $\dfrac{1.075 \text{ g soln/cm}^3}{1.0 \text{ g } H_2O/\text{cm}^3} \left| \dfrac{1.00 \text{ g } H_2O/\text{cm}^3}{} \right| \dfrac{1 \text{ lb}}{454 \text{ g}} \left| \dfrac{2.832 \times 10^4 \text{ cm}^3}{\text{ft}^3} \right. = 67.1 \dfrac{\textbf{lb soln}}{\textbf{ft}^3 \textbf{ soln}}$

c. $\dfrac{1.075 \text{ g soln/cm}^3}{1.0 \text{ g } H_2O/\text{cm}^3} \left| \dfrac{1.0 \text{ g } H_2O/\text{cm}^3}{} \right| \dfrac{12.7 \text{ g sugar}}{100 \text{ g soln}} \left| \dfrac{1000 \text{ cm}^3}{1 \text{ L}} \right. = 136 \dfrac{\textbf{g sugar}}{\textbf{L soln}}$

C ✓

Problem 1.4 A

A liquified mixture of n-butane, n-pentane and n-hexane has the following composition in percent.

n - C_4H_{10}	50	58
n - C_5H_{12}	30	72
n - C_6H_{14}	20	86

Calculate the weight fraction, mol fraction and mol percent of each component and also the average molecular weight of the mixture.

Solution

Note that the hydrocarbon mixture is liquid so that the composition is therefore in weight percent. It is convenient to use a weight basis and set up a table to make the calculations.

Basis: 100 kg

	% = kg	wt fr	MW		kg mol	mol fr
n - C_4H_{10}	50	**0.50**	58	50/58	0.86	**0.57**
n - C_5H_{12}	30	**0.30**	72	30/72	0.42	**0.28**
n - C_6H_{14}	20	**0.20**	86	20/86	0.23	**0.15**
	100	**1.00**			1.51	**1.00**

0.86/1.51

Average molecular weight $= \dfrac{\text{total mass}}{\text{total mol}} = \dfrac{100 \text{ kg}}{1.51 \text{ kg mol}} =$ **66**

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15 1.12