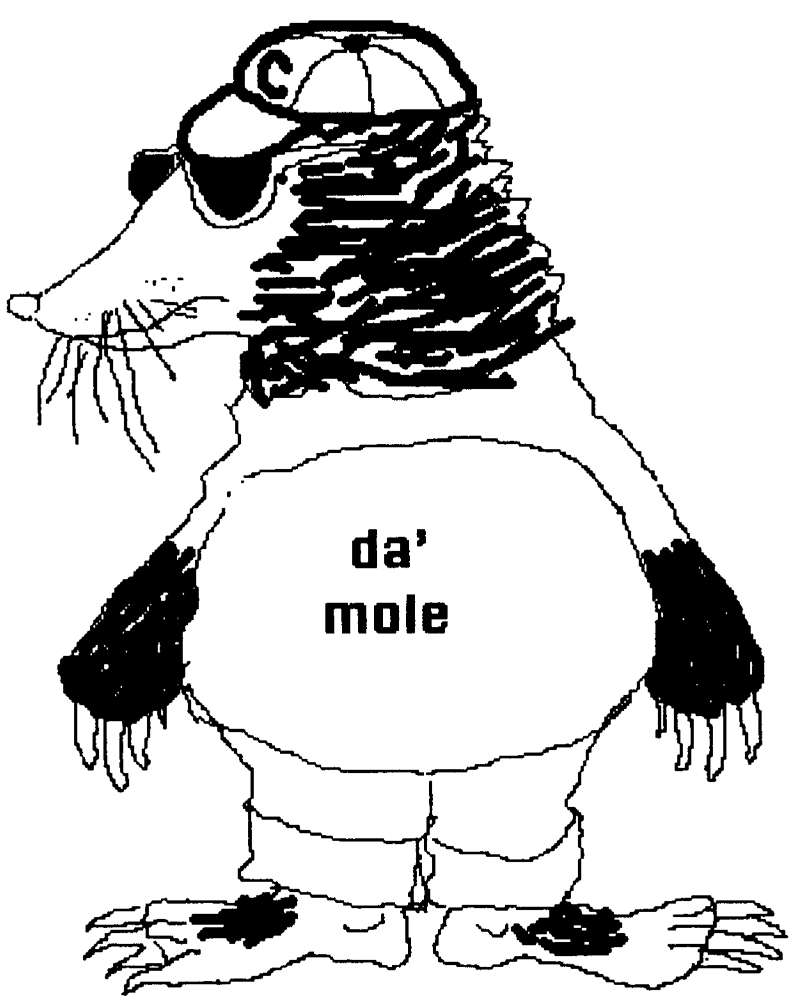
The Heart of Chemistry



Unit 4

Unit 4: The Mole ..... the Heart of Chemistry

I. The Significance of One Mole

Chemical symbols or formulas are shorthand representations of elements and compounds. In the simplest sense, they represent individual atoms or formula units(molecules). Since atoms are so very small, chemists usually deal with large numbers of atoms in the laboratory. The SI standard for the number of particles of a substance is the **mole**(abbreviated mol). By definition, **1 mole contains 6.02 x 10^{23} particles**(this number is called **Avogadro's number**, symbolized by N_A, in honor of a 19th century Italian scientist).

Chemists have found that 1 mole of atoms of any element **has a mass equal to the element's atomic mass** (sometimes called atomic weight) **expressed in grams.**

For example:

1 mole C = 6.02 x 10^{23} C atoms = 12.0111 g C

1 mole H = 6.02 x 10^{23} H atoms = 1.0079 g H

1 mole O = 6.02 x 10^{23} O atoms = 15.9994 g O

These relationships can be expressed as conversion factors, which allow chemists to count the number of particles present in a sample by massing it.

Sample Problem:
How many atoms of carbon are there in a sample that has a mass of 4.0 grams? How many moles of C are present?

$$\text{x atoms C} = 4.0 \text{ grams C} \left(\frac{6.02 \times 10^{23} \text{ C atoms}}{12.0111 \text{ grams C}}\right) = 2.0 \times 10^{23} \text{ C atoms}$$

$$\text{x moles C} = 4.0 \text{ grams C} \left(\frac{1 \text{ mole C}}{12.0111 \text{ grams C}}\right) = 0.33 \text{ mole C}$$

Problems:
Solve each of the following problems, using the appropriate conversion factors:

1. How many moles of oxygen atoms are contained in an 80.0 gram sample?

Ans: 5.00 moles O

2. How many atoms of zinc are contained in 15 moles of the metal?

Ans: 9.0 x 10^{24} atoms Zn

3. How many atoms of gold are contained in a 1.00 ounce ingot of gold? (Remember: 1 ounce = 28.35 grams)

Ans: 8.66 x 10^{22} atoms Au

4. What would be the mass of 0.125 moles of mercury atoms?

Ans: 25.1g Hg

5. What weighs more.......a mole of feathers or a mole of lead atoms? Explain your answer.

II. Determining the Molar Mass from a Chemical Formula

The **molar mass** of any chemical formula is calculated by adding the atomic masses of all of the atoms in the formula. For example, the chemical formula that represents the compound calcium phosphate is $Ca_3(PO_4)_2$. This formula shows that in one mole of this compound, there are 3 moles of Ca atoms that combine with 2 moles of the PO_4(phosphate) groups, which gives a total of 2 moles of P atoms and 8 moles of O atoms.

The total mass of each element in the compound is equal to the number of moles of atoms of the element multiplied by its atomic mass.

For $Ca_3(PO_4)_2$:

number of moles of atoms	x	atomic mass	=	total mass of element
3 moles of Ca	x	40.1 g Ca / mole	=	120.3 g Ca
2 moles of P	x	31.0 g P / mole	=	62.0 g P
8 moles of O	x	16.0 g O / mole	=	128.0 g O

The molar mass of calcium phosphate is equal to the sum of the total masses of each element

$$\text{molar mass} = 120.3\text{g Ca} + 62.0\text{g P} + 128.0\text{g O} = 310.3\text{g } Ca_3(PO_4)_2$$

Therefore, 1 mole $Ca_3(PO_4)_2$ = 6.02×10^{23} formula units $Ca_3(PO_4)_2$ = 310.3g

- Note that 1 mole of formula units of any compound still contains 6.02×10^{23} particles, but the mass is equal to the compound's molar mass (sometimes called molecular mass or formula mass) expressed in grams.

Molar mass, Avogadro's number of particles and the mole can also be used to create conversion factors when solving problems involving compounds.

Sample Problem:

How many formula units of $Ca_3(PO_4)_2$ are present in a sample that has a mass of 5.17 grams?

$$x \text{ formula units } Ca_3(PO_4)_2 = 5.17\text{g } Ca_3(PO_4)_2 \times \left(\frac{6.02 \times 10^{23} \text{ formula units}}{310.3 \text{ g } Ca_3(PO_4)_2}\right)$$

$$= 1.00 \times 10^{22} \text{ formula units } Ca_3(PO_4)_2$$

Problems:

Solve each of the following problems, using the appropriate conversion factors:

1. How many moles of TlBr contain 1.00×10^{26} formula units of TlBr?

Ans: 1.66×10^2 moles TlBr

2. How many molecules of H_2O are there in 0.400 mole of H_2O?

Ans: 2.41×10^{23} molecules H_2O

3. Find the number of moles of $CaBr_2$ that are equivalent to 76.0 grams $CaBr_2$.

Ans: 0.380 moles $CaBr_2$

4. Calculate the number of formula units of Ag_2Te in 46.0 grams of Ag_2Te.

Ans: 8.07×10^{22} units Ag_2Te

5. What is the mass of 1.00×10^{23} formula units of magnesium acetate?

Ans: 23.6g magnesium acetate

III. Mole, Mass and Volume Ratios Based on a Chemical Formula

The **Law of Definite Composition** states that any particular compound will always contain the same proportions of its elements. The definite composition can be described in terms of the number of atoms(moles) of each element present, the total masses of each element present, or in some cases, the volumes of gaseous elements that combined to form the compound. Since these ratios between the elements in a compound will always remain the same, they can also be used as conversion factors for problem solving.

For example, water has a chemical formula of H_2O. There are 2 atoms of hydrogen combined with every 1 atom of oxygen. In one mole of water molecules, there will be 2 moles of H atoms and 1 mole of O atoms. The total mass of each element(number of moles x atomic mass) present in 1 mole of water is 2 grams H and 16 grams O. Finally, whenever hydrogen gas reacts with oxygen gas to form water, they always combine in a "2 volumes H to 1 volume O" ratio. You may have noticed that the volume ratio of the gases is the same as the mole ratio of the elements. In 1811, Amadeo Avogadro explained this relationship by his hypothesis, "Equal volumes of gases at the same temperature and pressure contain equal numbers(moles) of particles." At standard conditions(STP) of 0°C and 1 atmosphere of air pressure, 1 mole of any gas will occupy a volume of 22.4 L. This volume is called the **molar volume** of a gas at STP, and can, of course, be used as another conversion factor in problem solving.

Sample Problem:
When a flame is brought near 10.0 grams of hydrogen gas inside of a balloon, it explodes as the hydrogen reacts with oxygen gas in the air. How many grams of water vapor will be formed?

$$x \text{ grams } H_2O = 10.0 \text{ grams H } \frac{(18 \text{ grams } H_2O)}{(2 \text{ grams H})} = 90 \text{ grams } H_2O$$

this is the mass ratio based on the formula

Problems:
Solve each of the following problems, using the appropriate conversion factors:

1. When copper metal is added to a solution containing 5.00 g $AgNO_3$ dissolved in it, the silver metal will precipitate out. How many grams of silver will form?

Ans: 3.18g Ag

2. In 1772, Joseph Priestley discovered laughing gas, N_2O. By the end of the century, a fad for the high society in England was to inhale laughing gas at small parties. How many cm^3 of the nitrogen gas would be needed to react completely with $2000cm^3$ of oxygen gas to produce the N_2O?

Ans: $4000cm^3$

3. If 0.25 mole of hydrated copper (II) sulfate, $CuSO_4 \bullet 5H_2O$ is heated in a drying oven overnight, how many grams of water will be lost?

Ans: 22.5g H_2O

4a. When potassium chlorate is heated in a test tube, it decomposes to form oxygen gas. What is the maximum mass of oxygen that could be formed by completely decomposing 2.00 g of potassium chlorate?

Ans: 0.783g O

4b. If the oxygen formed normally exists as diatomic molecules, O_2, approximately how many liters of gas are produced at STP?

Ans: 0.548 L O_2 at STP

IV. Determining the Percent Composition By Mass of a Compound

Since each compound always has a definite composition of its elements, the mass ratios can be converted into equivalent percentage values. The percentage of the total mass of a compound contributed by an element is the percent composition of that element in the compound.

Percent composition is calculated using the following formula:

$$\% \text{ element by mass} = \frac{\text{total mass of the element}}{\text{total mass of the compound}} \times 100$$

The masses of the elements and the compound can be determined from the chemical formula or from experimental data. Steps for calculating percent composition are listed below:

From a chemical formula:

1. Determine the total mass that an element contributes to the compound.

Example: For the compound copper (II) nitrate, $Cu(NO_3)_2$:

number of atoms(moles)	x	atomic mass	=	total mass of element
1 mole Cu	x	63.5 g Cu / mole	=	63.5 g Cu
2 moles N	x	14.0 g N / mole	=	28.0 g N
6 moles O	x	16.0 g O / mole	=	96.0 g O

2. Determine the total mass (molar mass) of the compound.
molar mass $Cu(NO_3)_2$ = 63.5 g Cu + 28.0 g N + 96.0 g O = 187.5 g $Cu(NO_3)_2$

3. Divide the total mass of the element by the total mass of the compound and multiply by 100(or use the % function on your calculator).

$$\% Cu = \frac{63.5 \text{ g Cu}}{187.5 \text{ g } Cu(NO_3)_2} \times 100 = 33.9\% \text{ Cu}$$

$$\% N = \frac{28.0 \text{ g N}}{187.5 \text{ g } Cu(NO_3)_2} \times 100 = 14.9\% \text{ N}$$

$$\% O = \frac{96.0 \text{ g O}}{187.5 \text{ g } Cu(NO_3)_2} \times 100 = 51.2\% \text{ O}$$

4. Check your work! The sum of the percentage compositions of all of the elements in the compound must add up to be 100%

From experimental data:
During an experiment, a pure copper penny with a mass of 3.42 g is dissolved in excess nitric acid and 10.17 g $Cu(NO_3)_2$ is formed. Calculate the percent composition by mass of copper in the $Cu(NO_3)_2$.

$$\% \text{ Cu} = \frac{3.42 \text{ g Cu}}{10.17 \text{ g } Cu(NO_3)_2} \times 100 = 33.6\% \text{ Cu}$$

•Note that any differences between the experimental percent composition and the theoretical percent composition based on the chemical formula are usually due to experimental error.

Problem:
Calculate the percent composition by mass of hydrogen and oxygen in water.

Ans: 11.1% H, 88.9% O

V. Determining Empirical and/or Molecular Formulas of Compounds

Empirical formulas express the smallest whole-number ratio of atoms in a chemical formula. It may not be the actual formula of the compound, but it is the simplest formula. The true formula will always be some whole number multiple of the empirical formula. For example, ionic compounds, such as NaCl, do not form simple molecules. A single granule of table salt may contain 2 billion Na^{1+} ions, but it will also contain 2 billion Cl^{1-} ions fused together in an alternating positive ion / negative ion pattern.

To determine the formula of an unknown compound given its percent composition or experimental mass proportions, use the following procedure:

1. Identify the elements and their percent composition or experimental masses.
2. Divide each element's % composition or experimental mass by its atomic mass to determine the relative number of moles of the element.
3. Determine the smallest ratio between the number of moles of all elements. This is done by dividing the relative number of moles of each element by the smallest mole value for any of the elements present.
4. Express this ratio as whole-numbers.

Sample problem 1:
A compound that contains 50.0% of each element by mass is formed while burning S in pure oxygen. What is the empirical formula of the compound?

Step 1. Determine the relative number of moles by $\frac{\% \text{ composition of element}}{\text{(atomic mass of element)}}$

$$\text{moles S} = \frac{50.0\% \text{ S}}{32\text{g S/mole}} = 1.56 \text{ moles S}$$

$$\text{moles O} = \frac{50.0\% \text{ O}}{16\text{g O/mole}} = 3.12 \text{ moles O}$$

Step 2. Determine the smallest ratio of moles

$$\text{moles S} = \frac{1.56 \text{ moles S}}{1.56 \text{ moles S}} = 1.00 \text{ mole S}$$

$$\text{moles O} = \frac{3.12 \text{ moles O}}{1.56 \text{ moles S}} = 2.00 \text{ moles O}$$

Step 3. Write this as a whole-number ratio of moles

The empirical formula = S_1O_2 or SO_2

Sample problem 2:
If the masses or moles of the reactants are given instead of percent compositions, then the actual mole ratios can be determined in step 1 of the procedure, rather than the relative number of moles. Otherwise, the procedure is the same. Example: 7.44 grams of a compound are formed while burning 2.74 grams of N in pure oxygen. What is the empirical formula of the compound?

Step 1. moles N = $\frac{2.74 \text{ g N}}{14 \text{ g N/mole}}$ = 0.196 mole N

moles O = $\frac{4.70 \text{ g O}}{16\text{g O/mole}}$ = 0.294 mole O (grams 0 = 7.44 g compound - 2.74 g N)

Step 2. moles N = $\frac{0.196 \text{ mole N}}{0.196 \text{ mole N}}$ = 1.00 mole N

moles O = $\frac{0.294 \text{ mole O}}{0.196 \text{ mole N}}$ = 1.50 mole O

Step 3. The empirical formula = $N_1O_{1.5}$ written as N_2O_3 using a whole-number ratio. The 1 to 1.5 mole ratio is doubled, rather than rounding off to N_1O_1 or N_1O_2.

Problems:

1. Determine the empirical formula of a compound that is 82.8% C and 17.2% H.

Ans: C_2H_5

2. A compound contains 52.6% Ni, 21.9% C and 25.5% N. What is the empirical formula of this compound?

Ans: NiC_2N_2

3. When 7.98 grams of iron rusts completely, 11.41 grams of the iron oxide is formed. Calculate the empirical formula of this rust.

Ans: Fe_2O_3

Compounds that are formed by the combination of two or more nonmetals will usually exist as individual molecules. When you determine the empirical formula of the compound, it might not accurately reflect the true nature of these molecules. For example, the empirical formula for glucose is CH_2O, but the compound actually exists as $C_6H_{12}O_6$ molecules. The **true molecular formula** of a molecular compound can be determined only if the actual molecular(molar) mass of the compound is known. The actual molar mass of the compound will always be some whole number multiple of the empirical formula's mass. The actual molar mass of glucose is 180g / mole.

To determine the true molecular formula of a compound, divide the true molecular mass given for the compound by the empirical formula mass.

$$\frac{\text{true molecular mass}}{\text{empirical formula mass}} = \text{a whole-number}$$

$$\frac{180 \text{ g of glucose / mole}}{30.0 \text{ g } CH_2O\text{ / mole}} = 6; \text{ the true molecular formula} = (CH_2O)_6 = C_6H_{12}O_6$$

VI. Determining the Formula of a Hydrated Compound

The chemical formulas of **hydrated solids**(that have water molecules physically trapped within their crystal structures) can be determined using the same steps described for determining an empirical formula. The ratio between the number of moles of the compound attached to the number of moles of water can be determined from percent composition or experimental mass values.

Sample problem:
In the lab, a 10.407 g sample of hydrated barium iodide is heated to drive off the water. The dry sample has a mass of 9.520g. What is the mole ratio between barium iodide and water? What is the formula of the hydrate?

• Mass of water lost by heating = 0.887g H_2O (10.407 g hydrated solid - 9.520 g dry solid)
• The chemical formula of barium iodide is BaI_2 and water is H_2O.
• Molar mass of BaI_2 = 391 g / mole.
[(1 mole Ba x 137 g Ba / mole) + (2 moles I x 127g I / mole)]
• Molar mass of H_2O = 18.0 g / mole
[(2 moles H x 1.0 g H / mole) + (1 mole O x 16.0 gO / mole)]

$$\text{moles } BaI_2 = 9.520 \text{ g } BaI_2 \times \left(\frac{1 \text{ mole } BaI_2}{391 \text{ g } BaI_2}\right) = 0.0243 \text{ mole } BaI_2$$

$$\text{moles } H_2O = 0.887 \text{ g } H_2O \times \left(\frac{1 \text{ mole } H_2O}{18.0 \text{ g } H_2O}\right) = 0.0493 \text{ mole } H_2O$$

Divide both the moles of BaI_2 and H_2O by 0.0243 moles to get 1 mole BaI_2 attached to 2 moles H_2O. The hydrate's formula is written as $BaI_2 \bullet 2H_2O$

Problems:

1. What is the true molecular formula of a compound that has an empirical formula CH_3 and a true molecular mass of 75 g/mole?

Ans: C_5H_{15}

2. Aspirin contains 60.0% C, 35.5% O and 4.48% H. It has a molecular mass of 180. g/mole. What is the molecular formula of aspirin?

Ans: $C_9O_4H_8$

3. When a hydrated compound is heated with a Bunsen burner, 0.763g of water are lost and 0.737g of magnesium sulfite remain in the beaker. Determine the formula of the hydrate.

Ans: $MgSO_3 \bullet 6H_2O$

4. What is the formula of a hydrated compound that contains 86.7% Mo_2S_5 and 13.3% H_2O?

Ans: $Mo_2S_5 \bullet 3H_2O$

VII. Molarity...... Concentration of Moles in Solution

Molarity is the ratio between the moles of a dissolved substance and the volume of the solution.

Mathematically, the molarity of a solution = $\dfrac{\text{\# of moles of solute(dissolved substance)}}{\text{number of liters of solution}}$

A one-molar(1 M) solution of hydrochloric acid contains one mole of HCl dissolved in one liter of solution. This means there are 36.5 grams of HCl dissolved in one liter of solution. These proportions of solute to solution will be the same, regardless of the volume of solution used. For example, in 500 mL of 1 M HCl(half of a liter), there will be 0.5 mole of HCl, or 18.25 g HCl.

Calculating Molarity of a Solution:

Dimensional analysis(factor-label method) is used to determine the molarity value of a solution, which will have the units of moles / liter.

Sample problem:
What is the molarity of 250 mL of solution containing 9.46g CsBr?

1. Calculate the moles of CsBr in 9.46g CsBr

$$\text{x moles CsBr} = \text{9.46g CsBr}\left(\frac{\text{1 mole CsBr}}{\text{213g CsBr}}\right) = \text{0.0444 mole CsBr}$$

2. Determine the liters of solution

$$\text{x liters soln} = \text{250mL soln}\left(\frac{\text{1 liter soln}}{\text{1000mL soln}}\right) = \text{0.250 liter soln}$$

3. Divide moles by liters to get molarity

$$\text{molarity} = \left(\frac{\text{0.0444 mole CsBr}}{\text{0.250 liter soln}}\right) = \text{0.178M CsBr}$$

Individual steps can be combined into one continuous problem, as shown below:

$$\text{molarity} = \frac{(\text{9.46 g CsBr})(\text{1 mole CsBr})(\text{1000mL soln})}{(\text{250mL soln}(\text{213g CsBr})(\text{1 liter soln})} = \text{0.178M CsBr}$$

Making a Solution:

Making a solution with a specific concentration involves determining the exact mass of solute needed to be dissolved in the required amount of water to produce the desired molarity.

Sample problem:
How would you make 500mL of 0.133M solution of $MnSeO_4$?

$$\text{x g } MnSeO_4 = \frac{\text{500mL soln}}{}\left(\frac{\text{1 liter}}{\text{1000mL}}\right)\left(\frac{0.133\text{mole } MnSeO_4}{\text{1 liter soln}}\right)\left(\frac{198\text{g } MnSeO_4}{\text{1 mole } MnSeO_4}\right)$$

Note the molarity is being used as a conversion factor. $= 13.2\text{g } MnSeO_4$

To make this solution, you would dissolve 13.2 grams $MnSeO_4$ in enough water to make 500mL of solution.

Problems:
Compute the molarity of the following solutions consisting of:

1. 49.9g $Pb(ClO_4)_2$ dissolved in 200. mL solution

Ans: 0.614 moles/L

2. 45.1g $CoSO_4$ dissolved in 250. mL solution

Ans: 1.16 moles/L

3. 41.3g $Fe(NO_3)_2$ dissolved in 100. mL solution

Ans: 2.30 moles/L

Describe how you would prepare each of the following solutions:

4. 250 mL of 4.00M RbOH solution

Ans: dissolve 102g RbOH in 250mL solution

5. 500. mL of 1.50M AgF solution

Ans: dissolve 95.2g AgF in 500mL solution

6. 250. mL of 0.00200M Tl_2CO_3 solution

Ans: dissolve 0.234g Tl_2CO_3 in 250mL solution

Diluting Concentrated Solutions:

When diluting a solution by adding more solvent, the amount of solute does not change. Since the total moles of solute remains constant during the dilution, the following relationship applies:

$$\textbf{Molarity}_{\text{initial}} \times \textbf{Volume}_{\text{initial}} = \textbf{Molarity}_{\text{final}} \times \textbf{Volume}_{\text{final}}$$

To prepare a certain volume of a solution with a specific molarity from a more concentrated solution, calculate the volume of the concentrate needed to produce the same number of moles of solute found in the given amount of diluted solution.

Sample problem:
How many mL of concentrated 12M HCl(aq) solution must be added to a sufficient amount of distilled water to produce 50.0 mL of 3.0M HCl(aq)?

$$\textbf{Molarity}_{\text{initial}} \times \textbf{Volume}_{\text{initial}} = \textbf{Molarity}_{\text{final}} \times \textbf{Volume}_{\text{final}}$$

$$\text{12M HCl(aq)} \times \textbf{Volume}_{\text{initial}} = \text{3.0M HCl(aq)} \times \text{50.0 mL}$$

$$\textbf{Volume}_{\text{initial}} = \frac{\text{3.0M HCl(aq)} \times \text{50.0 mL}}{\text{12M HCl(aq)}} = \text{12.5 mL of the 12M HCl(aq)}$$

Prepare the 50.0 mL of 3.0M HCl(aq) solution by slowly adding 12.5 mL of the 12M HCl(aq) to enough distilled water(approximately 37.5 mL of water) to make exactly 50.0 mL of the final solution. The amount of water may vary slightly, depending on how the acid and water bond together. **Caution:** Always add the concentrated acid to the water, never water to the acid, to dissipate any heat released by the dilution!

Problems:

1. How many mL of concentrated 18M H_2SO_4(aq) must be diluted with water to produce 25mL of 0.50M H_2SO_4(aq)?

Ans: 0.69mL 18M H_2SO_4(aq)

2. When 5.0 mL of concentrated nitric acid is diluted with enough water to produce 1.0 L of a solution, the final concentration is 0.079M. What is the original concentration of the nitric acid?

Ans: 16M

3. A chemical company produces vinegar by diluting concentrated acetic acid(that has a molarity of 17.4M) until the molarity equals 1.67M. If the business office just received an order for 50 liters of vinegar and the lab has only one 2.37 L bottle of concentrated acetic acid in stock, can they fill the order?

Ans: No, only 24.7 L of vinegar can be produced

VIII. Other Methods of Expressing Solution Concentration

Molality(m) is the ratio between the moles of solute and the kilograms of solvent used. Sometimes this expression of concentration is used instead of molarity because the proportion of solute to solvent is more consistent. The volume of solvent needed to make a solution with a specific molarity can vary, depending on the interaction between the solute and solvent particles.

$$\text{molality of a solution} = \frac{\text{number of moles of solute}}{\text{number of kilograms of solvent}}$$

Density is another common way of indicating concentration of a solution.

$$\text{density of a solution} = \frac{\text{grams of solution}}{\text{mL of solution}} = \frac{\text{grams of solute} + \text{grams of solvent}}{\text{mL of solution}}$$

Percent composition of the solution can be expressed as a volume or mass ratio between the amount of solute and the amount of solution.

$$\%\text{ by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100 \qquad \%\text{ by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

(mass solution = mass solute + mass solvent)

$$\%\text{ by mass per volume} = \frac{\text{grams of solute}}{\text{mL of solution}} \times 100$$

When percent composition is used to express concentration, it must be specified which of these three types of percent ratios are being represented.

The percentage composition value is an expression of the number of parts of solute per hundred parts of the mixture. For very dilute solutions, it is better to express the solute concentration as either **parts per million(ppm)** or **parts per billion(ppb)**. For example, 2-3 ppm of chlorine must be added to a swimming pool to kill harmful bacteria and chlorinate the pool. This is equivalent to 2-3 g of chlorine per 1,000,000 grams of water. Expressed as a mass percentage, this is only 0.0002% - 0.0003%. However, slight changes in this concentration can make a big difference in the mixture's properties, as you may have experienced when the water in the pool irritated your eyes.

The **mole fraction** represents the proportion between the moles of solute and the total moles of solute and solvent.

$$\text{mole fraction of a solute(or solvent)} = \frac{\text{moles of solute(or solvent)}}{\text{moles of solute + moles of solvent}}$$

Each of these methods for expressing concentration provides a specific comparison between the solute, solvent and/or solution. They can be used either individually or in combination as conversion factors during problem solving.

Sample problem 1:

Concentrated sulfuric acid has a density of 1.84g/mL and is 96% H_2SO_4 by mass. What is the molarity of this solution?

$$\frac{\text{x moles of } H_2SO_4}{\text{liter of solution}} = \left(\frac{1.84\text{g soln}}{1\text{mL of soln}}\right)\left(\frac{1000\text{ mL solution}}{1\text{ liter solution}}\right)\left(\frac{96\text{g } H_2SO_4}{100\text{g solution}}\right)\left(\frac{1\text{ mole } H_2SO_4}{98.1\text{g } H_2SO_4}\right)$$

molarity of solution = density of solution **x** metric conversion **x** % composition H_2SO_4 by mass **x** molar mass H_2SO_4

= 18.0 moles H_2SO_4 per liter solution or 18.0**M** $H_2SO_4(aq)$

Sample problem 2:

100-proof Scotch whisky contains 50% ethanol by volume. If the density of ethanol(C_2H_5OH) is 0.79 g/mL, then what is the molality of the solution?

Molality = number of moles ethanol per kg of water(the solvent)

$$= \left(\frac{0.79\text{g ethanol}}{1\text{mL ethanol}}\right)\left(\frac{1\text{ mole ethanol}}{46.0\text{g } C_2H_5OH}\right)\left(\frac{50\text{mL ethanol}}{100\text{mL soln}}\right)\left(\frac{100\text{mL soln}}{50\text{mL water}}\right)\left(\frac{1\text{mL } H_2O}{1\text{g } H_2O}\right)\left(\frac{1000\text{g } H_2O}{1\text{ kg } H_2O}\right)$$

= density ethanol **x** molar mass ethanol **x** % composition ethanol by volume **x** % composition H_2O by volume **x** density H_2O **x** metric conversion

= 17.2 moles ethanol per kg of water or 17.2 **m** $C_2H_5OH(aq)$

Problems:

1. A solution is prepared by dissolving 116g NaOH in water and then diluting to 250. mL. The density of this solution is 1.37 g/mL. Express the concentration of NaOH as:

a. percentage by mass

Ans: 33.9% NaOH

b. molarity

Ans: 11.6M NaOH

c. molality

Ans: 12.8m NaOH

d. mole fraction

Ans: 0.19 NaOH; 0.81 H_2O

2a. If you make a solution of grape juice from frozen concentrate by mixing 1 can of the concentrate with 3 cans of water, what is the percentage concentration by volume?

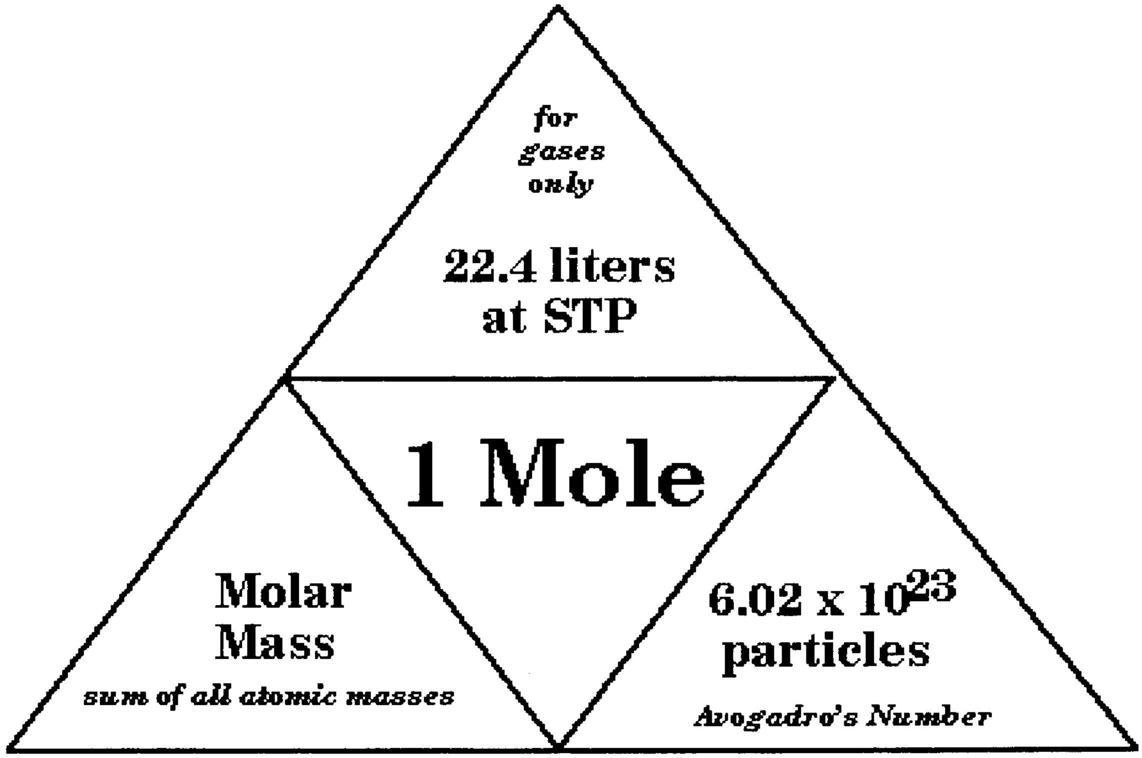
Ans: 25%

2b. If you remove 1.00 mL of the grape juice solution made in Part A above and dilute it with water to a volume of 1.000L in a volumetric flask, how many ppm of grape juice are in the new mixture?

Ans: 250 ppm

2c. How do you think the diluted grape juice in Part B above will differ from the original mixture made in Part A?

The mole triangle below illustrates the inter-relationships between the number of moles of a substance, its mass, the number of particles present, and the volume occupied at standard conditions, if the substance is a gas.



Unit 4 Objectives

Having studied the unit notes and done the problems, you should be able to:

1. Write formulas for chemical compounds given their name, name compounds from given chemical formulas and state the number of atoms of each element present in the compound. (These are skills acquired during Unit 3. Go back and review the Unit 3 study guide, if necessary.)
2. Use dimensional analysis to solve problems converting moles to mass to formula units.
3. Calculate the total mass of each element in a compound.
4. Calculate the molar mass of ionic and molecular formulas
5. Apply the mole, mass or volume ratios based on a chemical formula to determine the experimental proportions of the elements present in a compound.
6. Calculate the percent composition by mass of each element in a compound, using either the chemical formula or experimental data.
7. Find the empirical formula by determining the smallest whole-number ratio of atoms in a chemical formula.
8. Determine the true molecular formula of a molecular substance, given the actual molecular mass.
9. Determine the formula of a hydrated compound from either experimental data or the percent composition.
10. Compute the molarity of a solution given its composition.
11. Determine the amount of substance needed to prepare a solution with a definite molar concentration.