Lab. 3: Solution Concentrations

1. Molarity (in units of mol/L, molar, or M) or **molar concentration** denotes the number of moles of a given substance per liter of solution. Preparation of a solution of known molarity involves adding an accurately weighed amount of solute to a volumetric flask, adding some solvent to dissolve it, then adding more solvent to fill to the volume mark. M is used to abbreviate the units of mol/L.

The actual formula for **molarity** is:

$$\frac{\text{Moles of solute}}{\text{Liters of solution}} = \text{Molarity of solution.}$$

Such a solution may be described as "0.50 molar." It must be emphasized that a 0.5 molar solution contains 0.5 moles of solute in 1.0 liter of *solution*. This is *not* equivalent to 1.0 liter of **solvent**. A 0.5 mol/L solution will contain either slightly more or slightly less than 1 liter of solvent because the process of dissolution causes the volume of the liquid to increase or decrease.

For example: The molecular weight of sodium chloride (NaCl) is 58.44, so one gram molecular weight (= 1 mole) is 58.44g. If you dissolve 58.44g of NaCl in a final volume of 1 litre, you have made a 1M NaCl solution. To make a 0.1M Na Cl solution, you could weigh 5.844g of NaCl and dissolve it in 1 litre of water; OR 0.5844g of NaCl in 100mL of water.

M=Mole/L = m mole/mL ……....... (1)

Mole = wt.(gm)/m.wt = mg/m.wt ……..(2)

M = wt.(g)/m.wt /v/L …….(3)

M = Wt.(gm)/m.wt x 1000/V …......…(4)

Wt (gm) = M x m.wt x v/1000 …........(5)

Example:
What is the molarity of a solution made when water is added to 11 g $CaCl_2$ to make 100 mL of solution?

Solution:
11 g $CaCl_2$ / (110 g $CaCl_2$ / mol $CaCl_2$) = 0.10 mol $CaCl_2$
100 mL x 1 L / 1000 mL = 0.10 L
molarity = 0.10 mol / 0.10 L
molarity = 1.0 M

2. Molality (m): Molality is the number of moles of solute per kilogram of solvent. Because the density of water at 25°C is about 1 kilogram per liter, molality is approximately equal to molarity for dilute aqueous solutions at this temperature.

This is a useful approximation, but remember that it is only an approximation and doesn't apply when the solution is at a different temperature, it symbol to identify writing by the small word (m).

Determine mass of solvent in kg

Density = mass/volume   mass = density x volume

In Molality solution the final volume not equal to the one liter. Example/when we add the 58.5 gm of NaCl solute to the 1 litter of D.W the final volume become increasing to 1018ml.

Example:
How would you make 300 mL of a 0.5<u>M</u> NaOH solution?

First find molecular weights: Na = 23 0 = 16 H = 1 = NaOH = 40 (MW)

Then, 1<u>M</u> = 40 g/L

0.5<u>M</u> = 0.5 (40 g/L) = 20 g/L

But you only want 300 mL so . . .

20 /1000 mL = x /300 mL = 6 g = x, so we need 6 gram of NaoH to prepare the solution with 0.5 M

Question: A 4 g sugar cube (<u>Sucrose</u>) is dissolved in a 350 ml teacup of 80 °C water. What is the molality of the sugar solution?
Given: Density of water at 80° = 0.975 g/ml
Step 1 - Determine number of moles of sucrose in 4 g Solute is 4 g

$C_{12}H_{22}O_{11}$ = (12)(12) + (1)(22) + (16)(11)
$C_{12}H_{22}O_{11}$ = 144 + 22+ 176
$C_{12}H_{22}O_{11}$ = 342 g/mol
divide this amount into the size of the sample
4 g /(342 g/mol) = 0.0117 mol

Step 2 - Determine mass of solvent in kg.

Density = mass/volume
mass = density x volume
mass = 0.975 g/ml x 350 ml
mass = 341.25 g
Step 3 - Determine molality of the sugar solution.

molality = $mol_{solute} / m_{solvent}$
molality = 0.0117 mol / 0.341 kg
molality = 0.034 mol/kg the molality of the sugar solution.

3. Normality (N): Normality is equal to the *gram equivalent weight* of a solute per liter of solution. A gram equivalent weight or equivalent is a measure of the reactive capacity of a given molecule. Normality is the only concentration unit that is reaction dependent.

Example:
1 M sulfuric acid (H_2SO_4) is 2 N for acid-base reactions because each mole of surfuric acid provides 2 moles of H^+ ions. On the other hand, 1 M sulfuric acid is 1 N for sulfate precipitation, since 1 mole of sulfuric acid provides 1 mole of sulfate ions.
NOTE: This formula assumes that the solute has a percent assay=100% and that it is a solid. Remember, when actually making the solution, a small amount of solvent is used to dissolve the solute. Sufficient solvent is then added to bring the final volume to the specified amount. In other words, a sufficient quantity of solvent is added (q.s.) to yield the final volume, after the solute is first dissolved into a volume of solvent.

* equivalent weight of acid = m.wt/number of (H^+) atoms.
Ex: HCl = 36.5/1= 36.5
H_2SO_4 = 92/2 = 46

* equivalent weight of base =m.wt/number of (OH^-) atoms.
Ex: NaOH = 40/1= 40 , $Ca(OH)_2$=74/2=37

* equivalent weight of salt = m.wt/number of positive(+) or negative (-) charges.
Ex: H_3PO_4 = 100/3=33.33, NaCl 58.5/1=58.5, $CaCl_2$=111/2 = 55.5
We can able to use this formula to get weight of solute needed to solve in solution:
N = wt of solute in 1liter/eq.wt = >N=sp.gravity x (v%) x 1000/eq.wt
So the solute wt. requirement to dissolve in water requirement = eq.wt x (v) needed for preparation x asked concentration (N)/1000

N=Eq/L ……………… (1)

Eq = wt.(gm)/eq.wt ……..(2)

N= wt(g)/eq.wt /v/1000 …………(3)

N = wt(gm)/eq.wt x 1000/v ……(4)

Wt(gm)= N x eq.wt x v/1000 …. (5)

$\boxed{nN=M}$ …. N refer to the *equivalent number* ex:H_2SO_4 H_2 n = 2

$N_1V_1=N_2V_2$
Example : if you known that sp.gravity of HCl are 1.18, and the percentage rate of it equal to %34.46, and eq.wt of it are 36.5, find the normality of HCl require to preparation?
N= wt of solute in 1liter/eq.wt = > N=sp.gravity x (v%) x 1000/eq.wt
N= wt of solute in 1liter/eq.wt = > N= 1.18 x (34.46/100) x 1000/36.5 = 11.14

4.Percent solutions: There are three types of percent solutions. All are parts of solute per 100 total parts of solution. Based on the following definitions you may calculate the concentration of a solution or calculate how to make up a specific concentration.

1. **% W/W -** Percent of weight of solute in the total weight of the solution. The number of grams of solute in 100 grams of solution.

Example: A 100% (W/W) NaCl solution is made by weighing 100 g NaCL and dissolving in 100 g of solution.

2. **% W/V -** Percent of weight of solution in the total volume of solution. The number of grams of solute in 100 mL of solution. This is probably the least significant way of naming a solution, but the most common way of doing it.

Example: A 4% (W/V) NaCl solution is 4 g of NaCl in 100 mL of solution.

3. **% V/V -** Percent of volume of solute in the total volume of solution %V/V. Percent here is the number of milliliters of solute in 100 mL of solution.

Example:

A 10% (V/V) ethanol solution is 10 mL of ethanol in 100 mL of solution; unless otherwise stated, water is the solvent. So now, here are some applications

1. What is the percent concentration of a solution that you made by taking 5.85 g of NaCl and diluting to 100 mL with H_2O?

5.85 g/100 mL = 5.85% W/V solution of NaCl

2. What is the percent concentration of a solution that you made by taking 40 g of CaCl**2** and diluting to 500 mL with H**2**0?

You set up a proportion problem.

40 g /500 mL = Xg /100 mL = X = 8g = 8g/100 mL = 8% (W/V) solution

OR another way to look at this is, 40 grams solute is what percent of the 500 mL solution? The 100 is used to convert to percent.

3. How would you make 250 mL of a 8.5% NaCl solution? This works backwards from the others --

8.5%= 8.5 g /100 mL   Again, set up a proportion

8.5 g /100 mL= X / 250 mL   = 21.3 g = X

OR an alternative method is to say what is 8.5% of 250 mL?

250 x 0.085 = 21.3   Therefore you would need to weigh out 21.3 g NaCl and dilute to 250 mL with H**2**0.

4. How much (volume) 0.85% NaCl may be made from 2.55 g NaCl?

An 0.85% NaCl solution = 0.85 g/100 mL   Setting up a proportion again,

0.85 g /100 mL = 2.55 g/X = X = 300 mL Therefore, 300 mL of 0.85% NaCl may be made from 2.55 g NaCl.

5.Concentration: ppm and ppb: Parts per million (ppm) and parts per billion (ppb) are examples of expressing concentrations by mass. These units turn out to be convenient when the solute concentrations are very small (almost trace amounts).

1 : 1000 000 (1mg/L)   PPm   or 1 : 1000 000 000 PPb (1Mgm/L)

$$\frac{\text{mass solute}}{\text{mass solution}} \times 10^{6} = \text{concentration (ppm)}$$

$$\frac{\text{mass solute}}{\text{mass solution}} \times 10^{9} = \text{concentration (ppb)}$$

For example, suppose a 155.3 g sample of pond water is found to have $1.7\text{x}10^{-4}$ g of phosphates. What is the concentration of phosphates in ppm?

$$\frac{1.7\times10^{-4}\text{ g Phosphates}}{1.553\times10^{2}\text{ g Solution}} \times 10^{6} = 1.1\text{ ppm}$$

A similar procedure would be followed to calculate ppb. In the above example the pond water would be 1,100 ppb.

Now suppose we have 400 g sample of pond water and it has a concentration of 3.5 ppm dissolved nitrates. What is the mass of dissolved nitrates in this sample?

$$400\text{ g }\cancel{\text{solution}} \times \frac{3.5\times10^{-6}\text{ g nitrates}}{1\text{ g of }\cancel{\text{solution}}} = 1.4\text{ mg nitrates}$$