# Units

1. Units to describe a solution

   a. Molarity

   b. Molality

   c. Mole Fraction

   d. Mass Percent

2. What are the units of molarity?

$$\frac{\text{moles of solute}}{\text{L of sol'n}} = \text{Molarity (M)}$$

3. What are the units of molality?

$$\text{Molality (m)} = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

4. How do you determine the mole fraction?

$$\text{Mass \%} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

5. How do you determine the mass percent?

$$\text{Mole Fraction } (\chi_a) = \frac{\text{moles}_a}{\text{moles}_{total}}$$

6. Calculate the molar concentration of all the ions in the following *strong* electrolytes.

a. 75g of KCl dissolved in 250 mL of water.

$$75\text{ g KCl} \times \frac{1.0\text{ mol KCl}}{74.55\text{g KCl}} = 1.0\text{ mol KCl}$$

$$\frac{1.0\text{ mol KCl}}{250 \times 10^{3}\text{ L sol'n}} = 4.0\text{ M KCl}$$

$$K^{+} \qquad Cl^{-}$$

It is important to note that the solution is not one made up of KCl molecules, but is a solution made up of $K^+$ and $Cl^-$ ions. That is why the question asked for the concentration for the *ions* in solution. That means that we are not concerned with the concentration of KCl as whole, but as individual components.

$$\frac{4.0\text{ mol }\cancel{\text{KCL}}}{\text{L of sol'n}} \times \frac{1\text{ mol K}^{+}}{1\text{ mol }\cancel{\text{KCL}}} = 4.0\text{ M K}^{+}$$

$$\frac{4.0\text{ mol }\cancel{\text{KCL}}}{\text{L of sol'n}} \times \frac{1\text{ mol Cl}^{-}}{1\text{ mol }\cancel{\text{KCL}}} = 4.0\text{ M Cl}^{-}$$

b. 100.0g of $MgF_2$ dissolved in 500.0 mL of water

$$100.0\text{ g MgF}_2 \times \frac{1\text{ mol MgF}_2}{62.31\text{ g MgF}_2} = 1.605\text{ mol MgF}_2$$

$$\frac{1.605\text{ mol MgF}_2}{0.500\text{L sol'n}} = 3.21\text{M MgF}_2$$

$$Mg^{2+} \qquad 2F^{-}$$

Now we solve for the concetrations of the individual ions in solution, just like the previous example.

$$\frac{3.21\ \cancel{\text{mol MgF}_2}}{\text{L of sol'n}} \times \frac{1\ \text{mol Mg}^{2+}}{1\ \cancel{\text{mol MgF}_2}} = 3.21\ \text{M Mg}^{2+}$$

$$\frac{3.21\ \cancel{\text{mol MgF}_2}}{\text{L of sol'n}} \times \frac{2\ \text{mole F}^-}{1\ \cancel{\text{mol MgF}_2}} = 6.42\ \text{M F}^-$$

7. An aqueous antifreeze solution is 40.0% ethylene glycol ($C_2H_6O_2$) by mass. The density of the solution is 1.05 g/cm$^3$. Calculate the molality, molarity, and mole fraction of ethylene glycol.

In this case the solute is the ethylene glycol and the solvent is water.

$$\text{Molality (m)} = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

We can use the percent by mass of the ethylene glycol o determine the moles of ethylene glycol and kg of solvent.

Remember that mass percent is:

$$\text{Mass \%} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

For ease, we will assume that we have 100 g of solution. This makes our lives easier because this means that we have 40.0 g of ethylene glycol.

The important part to this question is realizing that in molality the denominator is terms of solvent *only* and in mass% the denominator is in terms of the *whole* solution.

This means that if we assume a 100g solution and 40g of it were solute. The remaining 60g are solvent. Leaving us with the following fraction:

$$\frac{40.0\text{ g } C_2H_6O_2}{60\text{g } H_2O} \; \frac{1000\text{g}}{\text{kg}} \; \frac{1\text{ mol } C_2H_6O_2}{62.08\text{ g}} = \boxed{\frac{10.7\text{ mol } C_2H_6O_2}{\text{kg } H_2O}}$$

Now we turn to molarity:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{L of sol'n}}$$

I will start with the same initial point as for molality. The key difference here is that in molarity, the denominator is terms of the entire solution.

$$\frac{40.0\text{ g } C_2H_6O_2}{100\text{g sol'n}} \; \frac{1.05\text{ g}}{\text{cm}^3} \; \frac{1000\text{ cm}^3}{\text{L}} \; \frac{1\text{ mol } C_2H_6O_2}{62.08\text{ g}} = \boxed{\frac{6.77\text{ mol } C_2H_6O_2}{\text{L sol'n}}}$$

Lastly we will solve for the mole fraction:

$$\text{Mole Fraction } (\chi_a) = \frac{\text{moles}_a}{\text{moles}_{total}}$$

In this case we will use the percent by mass information to solve for the moles of ethylene glycol and the moles of water. Our assumption will still be that we are dealing with 100g of solution.

$$40.0\text{ g } C_2H_6O_2 \; \frac{1\text{ mol } C_2H_6O_2}{62.08\text{ g}} = 0.644\text{ moles } C_2H_6O_2$$

$$60.0\text{ g } H_2O \; \frac{1\text{ mol } H_2O}{18.02\text{ g}} = 3.33\text{ mol } H_2O$$

Now take this information and plug into the mole fraction equation:

$$\frac{0.644\text{ mol } C_2H_6O_2}{0.644\text{ mol } C_2H_6O_2 + 3.33\text{ mol } H_2O} = \boxed{0.162}$$

8. What is a dilution?

In a dilution the volume of a solution is increased but the moles of solute remains the same.

9. What equation is useful in a dilution?

$$\mathbf{M_1V_1 = M_2V_2}$$

$M_1$ = Initial molarity
$V_1$ = Initial sol'n volume

$M_2$ = Final molarity
$V_2$ = Final sol'n volume

*Caution – Only use this equation when a substance is NOT undergoing a reaction.

10. How much water must be added to a 2.00M stock sol'n of NaOH to make 100.0 mL of a 0.100M NaOH sol'n?

We know that this is a dilution question because we are told that all we are doing is adding water to the solution. Adding water will increase the volume of solution, but not affect the moles of solute (NaOH) present in solution. This means that it is completely fine to use $M_1V_1 = M_2V_2$

(100.0 mL) (0.100 M) = ($V_2$) (2.00M)
$V_2$ =5.00 mL of **stock** solution needed

It is important to note that the question is asking for the amount of water required – we just solved for the amount of the concentrated stock solution we would need to dilute down.

Amount of water that must be added is
100.0 mL – 5.00 mL = **95.0 mL water**