<u>Key Worksheet 1: Solution Composition Worksheet</u>

Objectives: To be able to define, calculate, and use the following units of concentration:

Molarity: $[\mathrm{X}] = \dfrac{\text{moles X}}{\text{Liters solution}}$

Mass %: $\text{Mass \% X} = \dfrac{\text{Mass of X}}{\text{Total mass of solution}} \times 100\%$ or $\dfrac{\text{Grams of X}}{\text{100 grams of solution}}$

Molality: $m_{\mathrm{x}} = \dfrac{\text{moles x}}{\text{kg solvent}}$

Normality: $\mathrm{N} = \dfrac{\text{Equivalents (Eq)}}{\text{Liters solution}}$

For an acid one equivalent (Eq) is defined as the number of moles of protons one mole of the acid can donate (the moles of acidic protons).

For a base one equivalent (Eq) is defined as the number of moles of protons one mole of the base can accepts. For bases that are hydroxides this is the moles of hydroxide in one mole of the base.

For an oxidizing agent in a redox reaction one equivalent (Eq) is defined the number of moles of electrons one mole of the oxidizing agent can accept.

For an reducing agent in a redox reaction one equivalent (Eq) is defined the number of moles of electrons one mole of the reducing agent can donate.

Mole Fraction: $\chi_{\mathrm{x}} = \dfrac{\text{Moles X}}{\text{Total moles in solution}}$

1.) 27.2 grams of phosphoric acid, H_3PO_4, is dissolved in enough water to make 175 mL of solution. Calculate the molarity of H_3PO_4.

$$\frac{27.2\text{ g }H_3PO_4}{97.994\text{ g/mol}} = 0.277_5\text{ mol }H_3PO_4$$

$$\frac{0.277_5\text{ mol }H_3PO_4}{0.175\text{ L solution}} = 1.58_6\text{ M}$$

Molarity = 1.59 M

2.) 9.34 grams of $CaCl_2$ are dissolved in 149 grams of water. Calculate the mass percent of calcium chloride in the resulting solution.

$$\frac{9.34\text{ g }CaCl_2}{(9.34\text{ g }CaCl_2 + 149\text{ g }H_2O)} \times 100\% = 5.89_8\%$$

Mass % = 5.90%

3.) 17.32 grams of KBr is dissolved in 249 grams of water. Assuming the potassium bromide completely dissociates, calculate the molality of ions in the resulting solution.

$$\frac{17.32\text{ g KBr}}{119.002\text{ g/mol}} = 0.1455_4\text{ mol KBr} \qquad 0.1455_4\text{ mol KBr} \times \frac{2\text{ moles ions}}{1\text{ mol KBr}} = 0.2910_8\text{ moles ion}$$

$$\frac{0.2910_8\text{ mol ions}}{0.249\text{ kg }H_2O} = 1.16_9\text{ m}$$

Molality = 1.17 m

4.) 79.44 grams of H_2SO_4 is dissolved in enough water to make 275 mL of solution. Calculate the normality of sulfuric acid in the resulting solution.

Note that there are 2 moles of acidic protons for every one mole of H_2SO_4.

$$\frac{79.44\text{ g }H_2SO_4}{98.078\text{ g/mol}} = 0.8099_6\text{ mol }H_2SO_4 \quad \frac{0.8099_6\text{ mol }H_2SO_4}{0.275\text{ L solution}} = 2.94_5\text{ M}$$

$$2.94_5\text{ M} \times 2\text{ mol }H^+ = 5.89_0\text{ N}$$

Normality = 5.89 N

5.) 5.25 pounds of dextrose ($C_6H_{12}O_6$) is dissolved in 11.3 gallons of water. (1.000 lb = 453.5 g, 1.000 gal. = 3.785 L, take the density of water to be 1.00 g/mL) Calculate the mole fraction of dextrose in the resulting solution.

$$11.3 \text{ gal. } \mathrm{H_2O} \times \frac{3.785 \text{ L}}{1 \text{ gal.}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1.00 \text{ g } \mathrm{H_2O}}{1 \text{ mL } \mathrm{H_2O}} = 4.27_7 \times 10^4 \text{ g } \mathrm{H_2O}$$

$$\frac{4.27_7 \times 10^4 \text{ g } \mathrm{H_2O}}{18.015 \text{ g/mol}} = 2.37_4 \times 10^3 \text{ mol } \mathrm{H_2O}$$

$$5.25 \text{ lb. dextrose} \times \frac{453.59 \text{ g}}{1 \text{ lb.}} \times \frac{1 \text{ mol dextrose}}{180.150 \text{ g dextrose}} = 13.2_1 \text{ mol dextrose}$$

$$\chi_{\text{dextrose}} = \frac{13.2_1 \text{ mol dextrose}}{13.2_1 \text{ mol dextrose} + 2.37_4 \times 10^3 \text{ mol } \mathrm{H_2O}} = 0.00553_6$$

Mole Fraction = <u>0.00554</u>

6.) Concentrated nitric acid is 70.0% HNO_3 by mass, which is 15.9 M. Calculate the density and molality of concentrated nitric acid.
a) Assume exactly 1 liter of solution. That means we have 15.9 moles of HNO_3.

Mass of HNO_3 is then: $15.9 \text{ mol } \mathrm{HNO_3} \times \frac{63.012 \text{ g } \mathrm{HNO_3}}{1 \text{ mol } \mathrm{HNO_3}} = 1.00_1 \times 10^3 \text{ g } \mathrm{HNO_3}$

We can now find the mass of the solution:

$$1.00_1 \times 10^3 \text{ g } \mathrm{HNO_3} \times \frac{100 \text{ g solution}}{70.0 \text{ g } \mathrm{HNO_3}} = 1.43_1 \times 10^3 \text{ g solution}$$

$$d_{\text{solution}} = \frac{\text{Mass of solution}}{\text{volume of solution}} = \frac{1.43_1 \times 10^3 \text{ g solution}}{1000 \text{ mL solution}} = 1.43_1 \text{ g/mL}$$

Density = <u>1.43 g/mL</u>

b)

$$\text{molality} = \frac{\text{mol } \mathrm{HNO_3}}{\text{kg } \mathrm{H_2O}}$$

Mass of H_2O = $1.43_1 x 10^3$ g solution – $1.00_1 x 10^3$ g HNO_3 = $4.2_{93} x 10^2$ g H_2O = 0.42_{93} kg H_2O

$$\text{molality} = \frac{15.9 \text{ mol } \mathrm{HNO_3}}{0.42_{93} \text{ kg } \mathrm{H_2O}} = 37._0 \ m$$

Molality = <u>37 m</u>

7.) Suppose you are brewing a batch of beer. You need to have a total of 101 ppm of Ca^{2+} in 12.0 gallons of water for the mash. You must have equal parts of $CaCl_2$ and $CaSO_4$ in the water to accomplish this. How many grams of each should you add? What is the molarity of Ca^{2+}, $CaCl_2$, and $CaSO_4$ in the resulting solution? (Ca^{2+} = 40.078 g/mol, $CaCl_2$ = 110.984 g/mol, $CaSO_4$ = 136.140 g/mol, 1.000 gal. = 3.785 L)

a)

$$12.0\text{ gal. solution} \times \frac{3.785\text{ L}}{1\text{ gal.}} \times \frac{1.00 \times 10^3\text{g solution}}{1\text{ L solution}} = 4.54_2 \times 10^4\text{ g solution}$$

$$4.54_2 \times 10^4\text{ g solution} \times \frac{101\text{ g Ca}^{2+}}{1 \times 10^6\text{ g solution}} = 4.58_7\text{ g Ca}^{2+}$$

Let x = mass of $CaCl_2$ and y = mass of $CaSO_4$.

$$4.58_7\text{ g Ca}^{2+} = x\text{ g CaCl}_2 \times \frac{40.078\text{ g Ca}^{2+}}{110.984\text{ g CaCl}_2} + y\text{ g CaSO}_4 \times \frac{40.078\text{ g Ca}^{2+}}{136.140\text{ g CaSO}_4}$$

Since we need equal parts of each compound, $x = y$. This gives us:

$$4.58_7\text{ g Ca}^{2+} = x\text{ g CaCl}_2 \times \frac{40.078\text{ g Ca}^{2+}}{110.984\text{ g CaCl}_2} + x\text{ g CaSO}_4 \times \frac{40.078\text{ g Ca}^{2+}}{136.140\text{ g CaSO}_4}$$

$$\Rightarrow x = 6.99_8\text{ g}$$

mass $CaCl_2$ = <u>7.00 g</u>

b)

mass $CaSO_4$ = <u>7.00 g</u>

c)

$$\frac{4.58_7\text{ g Ca}^{2+}}{40.078\text{ g/mol}} = 0.114_4\text{ mol Ca}^{2+} \qquad 12.0\text{ gal. solution} \times \frac{3.785\text{ L}}{1\text{ gal.}} = 45.4_2\text{ L solution}$$

$$\left[\text{Ca}^{2+}\right] = \frac{0.114_4\text{ mol Ca}^{2+}}{45.4_2\text{ L solution}} = 0.00252_0\ M$$

[Ca^{2+}] = <u>0.00252 M</u>

d)

$$\frac{7.00\text{ g CaCl}_2}{110.984\text{ g/mol}} = 0.0630_7\text{ mol CaCl}_2 \qquad \frac{0.0630_7\text{ mol CaCl}_2}{45.4_2\text{ L solution}} = 0.00138_8\frac{\text{mol CaCl}_2}{\text{L solution}}$$

[$CaCl_2$] = <u>0.00139 M</u>

e)

$$\frac{7.00\text{ g CaSO}_4}{136.140\text{ g/mol}} = 0.0514_1\text{ mol CaSO}_4 \qquad \frac{0.0514_1\text{ mol CaSO}_4}{45.4_2\text{ L solution}} = 0.00113_2\frac{\text{mol CaSO}_4}{\text{L solution}}$$

[$CaSO_4$] = <u>0.00113 M</u>

8.) Find the mass percent of $FeCl_3$ in a solution whose density is 1.42 g/mL and whose molarity is 1.19 M.

Assume we have exactly 1 liter of solution.

$$1.19 \text{ mol FeCl}_3 \times \frac{162.204 \text{ g FeCl}_3}{1 \text{ mol FeCl}_3} = 193._0 \text{ g FeCl}_3$$

$$1000 \text{ mL solution} \times \frac{1.42 \text{ g solution}}{1 \text{ mL solution}} = 1,420 \text{ g solution}$$

$$\frac{193._0 \text{ g FeCl}_3}{1,420 \text{ g solution}} \times 100\% = 13.5_9\%$$

Mass% $FeCl_3$ = <u>13.6%</u>

9.) Calculate the normality of Cu^{2+} in a solution that is made by dissolving 8.77 g of $Au(NO_3)_3$ in enough water to make 175 mL of solution when copper metal is added to the solution. The relevant reaction is

$$3\text{Cu}(s) + 2\text{Au}^{3+}(aq) \rightarrow 2\text{Au}(s) + 3\text{Cu}^{2+}(aq)$$

$$\frac{8.77 \text{ g Au(NO}_3)_3}{382.979 \text{ g/mol}} = 0.0228_9 \text{ mol Au(NO}_3)_3 \times \frac{1 \text{ mol Au}^{3+}}{1 \text{ mol Au(NO}_3)_3} = 0.0228_9 \text{ mol Au}^{3+}$$

$$\frac{0.0228_9 \text{ mol Au}^{3+}}{0.175 \text{ L}} \times \frac{3 \text{ mol Cu}^{2+}}{2 \text{ mol Au}^{3+}} \times \frac{2 \text{ equivalent}}{1 \text{ mol Cu}^{2+}} = 0.392_5 \; N$$

Normality = <u>0.393 N</u>