Chemistry 6A F2007

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Friday

11/9/07

Announcement:

This weeks experiment (Atomic Spectra/Flame Test) is due next week, even though there is no lab scheduled for the next two weeks.

Monday's Lab must turn in the lab by Tuesday (11/13)
Tuesday's Lab must turn in the lab by Tuesday (11/13)
Thursday's Lab must turn in the lab by Thursday (11/15)

Late labs will have points deducted.

The relative amounts of ***solute*** in the solution is expressed by a ***concentration***.

Molarity:

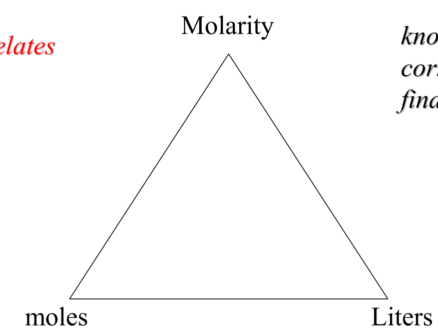
The ratio of moles of solvent to liters of solute

$$\text{molarity (M)} = \frac{\text{moles of solute}}{\text{L of solution}}$$

Moles/Liters and Molarity:

Molarity relates mols and volume (L)

knowing 2 corners, you can find the 3rd



If you know moles & L, you know Molarity (M) if you know Molarity and volume, you know moles!

Molarity × Volume = moles

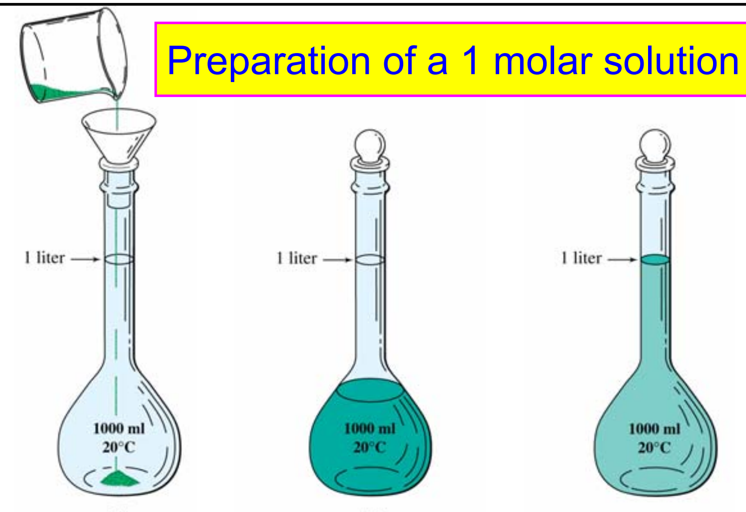
$$\frac{\text{mols}}{\text{L}} \times \text{L} = \text{moles}$$

if you know mols and molarity, you know volume!

$$\text{moles} \times \frac{1}{\text{M}} = \text{Volume}$$

$$\text{mol} \times \frac{\text{L}}{\text{mol}} = \text{L}$$

Preparation of a 1 molar solution



(a) Add 1 mole of solute to a 1-liter volumetric flask

(b) Dissolve in solvent

(c) Add solvent to the 1-liter mark and mix thoroughly

Making up solutions of *known* concentration:

To make "stock" solution of known molarity, one must:

1. Add a carefully measured amt. of solute to a ***volumetric flask***.

A volumetric flask is a piece of laboratory glassware that is accurately calibrated to a precise know volume at a known temperature.

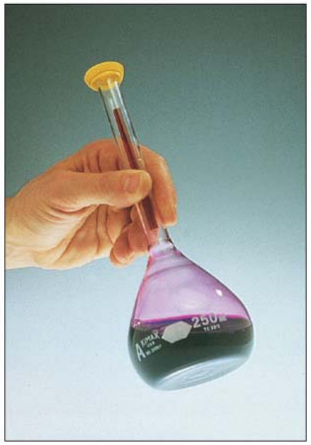
Making up solutions of *known* concentration:

To make "stock" solution of known molarity, one must:

1. Add a carefully measured amt. of solute to a ***volumetric flask***.

2. Fill the volumetric flask partially with the solvent to dissolve the solute.

One must be sure that the solute (if it is in solid form) is completely dissolved before the addition of more solvent.



Making up solutions of *known* concentration:

To make "stock" solution of known molarity, one must:

1. Add a carefully measured amt. of solute to a ***volumetric flask***.
2. Fill the volumetric flask partially with the solvent to dissolve the solute.
3. Fill the volumetric to the calibration mark using a bottle, then a dropper.

The bottom of the curved portion of the meniscus must be even with the calibration mark.

A mark on the neck of a volumetric flask indicates a volume of exactly 250 mL at 25 °C.

Making up solutions of *known* concentration:

Knowing the volume of the flask and the moles of solute, one can determine the molarity of the solution!

A mark on the neck of a volumetric flask indicates a volume of exactly 250 mL at 25 °C.

A student adds 25.15 g of sodium sulfide into a 500.0mL volumetric flask then fills the solution to the calibration mark with water. What is the molarity of this solution.

$$\text{molarity (M)} = \frac{\text{moles of solute}}{\text{L of solution}}$$

$$\text{g } Na_2S \longrightarrow \text{mols } Na_2S \longrightarrow M(Na_2S)$$

$$M(Na_2S) = \frac{25.15\text{g } Na_2S \times \frac{1 \text{ mol } Na_2S}{78.05\text{g } Na_2S}}{500.0\text{mL} \times \frac{1\text{ L}}{10^3\text{ mL}}} = 0.6445\text{M } Na_2S$$

How many grams of sodium sulfide are there in 25.1mL of a 0.6445M sodium sulfide solution?

$$\text{Volume (L)} \times \text{molarity(mol/L)} = \text{moles}$$

$$\text{moles} \times \text{molar mass} = \text{grams}$$

$$\text{mL} \longrightarrow \text{L} \longrightarrow \text{moles} \longrightarrow \text{grams}$$

$$25.1\text{mL} \times \frac{1\text{ L}}{10^3\text{mL}} \times \frac{0.6445\text{mol } Na_2S}{1\text{ L}} \times \frac{78.05\text{g } Na_2S}{\text{mol } Na_2S}$$

$$= 1.26\text{g } Na_2S$$

What is the sodium ion concentration of this solution?

$$Na_2S(aq) \longrightarrow 2Na^+(aq) + S^{2-}(aq)$$

When the salt dissolves in solution, 2 moles of sodium ion result for every one mole of the sodium sulfide salt.

$$0.6445\text{M } \cancel{Na_2S} \times \frac{2\text{mol } Na^+}{1\text{mol } \cancel{Na_2S}} = 1.289\text{M } Na^+$$

The concentration of ions depends upon the molar ratios in the salt.

How many moles of sodium ion are there in 25.1mL of a 0.6445M sodium sulfide solution?

$$\text{Volume (L)} \times \text{molarity(mol/L)} = \text{moles}$$

The moles of Na^+ are found by:

$$25.1\text{mL} \times \frac{1\text{ L}}{10^3\text{mL}} \times \frac{0.6445\text{mol } Na_2S}{1\text{ L}} \times \frac{2\text{mol } Na^+}{1\text{mol } Na_2S} =$$

0.0324 mols Na^+

Suppose one wants to prepare 250.0 mL of a 0.105 M solution of $AgNO_3$. How would this be done?

$$\text{Volume (L)} \times \text{molarity(mol/L)} = \text{moles}$$

$$\text{moles} \times \text{molar mass} = \text{grams}$$

volume ⟶ mols ⟶ grams *(that need to be added to 0.250 L)*

$$250.0\cancel{\text{mL}} \times \frac{\cancel{1\text{ L}}}{10^3\cancel{\text{mL}}} \times \frac{0.105\cancel{\text{mol}}\ AgNO_3}{\cancel{1\text{ L}}} \times \frac{169.88\text{g } AgNO_3}{\cancel{1\text{ mol}}}$$

$$= 4.46\text{g } AgNO_3$$

add 4.46g of $AgNO_3$ to a 250.0mL volumetric flask and dilute to the calibration mark with water.

A 5.37 % solution of potassium nitrate has a density of 1.11 g/mL. Calculate the molarity of the solution.

$$5.37\% = \frac{5.37\text{ g } KNO_3}{100.0\text{ g solution}} \qquad d = \frac{1.11\text{ g solution}}{1.00\text{ mL solution}}$$

$$\frac{5.37\ \cancel{\text{g } KNO_3}}{100.0\ \cancel{\text{g solution}}} \times \frac{1.11\ \cancel{\text{g solution}}}{1.00\ \cancel{\text{mL solution}}} \times \frac{10^3\ \cancel{\text{mL}}}{1\text{L}} \times \frac{1\text{ mol } KNO_3}{101.11\ \cancel{\text{g } KNO_3}} =$$

$$= 0.590\text{ mol } KNO_3/\text{L}$$

Solution Stoichiometry:

When solutions mix, a chemical reaction may result. One can calculate the concentrations of products and reactants based on the stoichiometry of the reaction.

Consider solutions of nitric acid and potassium hydroxide:

$HNO_3(aq)$ and $KOH(aq)$

When they mix: acid + base make a salt + water

$$HNO_3(aq) + KOH(aq) \longrightarrow KNO_3(aq) + H_2O(l)$$

How many mL of 0.125M nitric acid are needed to completely neutralize 25.1mL of 0.105M potassium hydroxide?

$$HNO_3(aq) + KOH(aq) \longrightarrow KNO_3(aq) + H_2O(l)$$

$$\text{mL KOH} \longrightarrow \text{mols KOH} \longrightarrow \text{mols } HNO_3 \longrightarrow \text{mL } HNO_3$$

$$25.1\text{mL} \times \frac{1\text{ L}}{10^3\text{mL}} \times \frac{0.105\text{mol KOH}}{1\text{ L}} \times \frac{1\text{mol } HNO_3}{1\text{mol KOH}} \times \frac{1\text{ L}}{0.125\text{mol } HNO_3} \times \frac{10^3\text{mL}}{1\text{L}} = 21.1\text{mL}$$

How many grams of silver chloride will result form the mixing of 25.1 mL of 0.125M NaCl and 15.6 mL of 0.105M $AgNO_3$?

Step 1: Write the balanced chemical reaction.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Step 2: Recognize that moles can be found form Molarity and volume

$$\text{Volume (L)} \times \text{molarity(mol/L)} = \text{moles}$$

Step 3: Recognize that this is a ***limiting reactant problem***;
The ***LR*** will determine the maximum amount of product that can form.

How many grams of silver chloride will result form the mixing of 25.1 mL of 0.125M NaCl and 15.6 mL of 0.105M $AgNO_3$?

$$25.1\text{mL} \times \frac{\text{L}}{10^3\text{mL}} \times \frac{0.125\text{mol NaCl}}{\text{L}} \times \frac{1\text{ mol AgCl}}{1\text{ mol NaCl}} \times \frac{143.4\text{g AgCl}}{\text{mol AgCl}}$$

~~= 0.450g AgCl~~

$$15.6\text{mL} \times \frac{\text{L}}{10^3\text{mL}} \times \frac{0.105\text{mol } AgNO_3}{\text{L}} \times \frac{1\text{ mol AgCl}}{1\text{ mol } AgNO_3} \times \frac{143.4\text{g AgCl}}{\text{mol AgCl}}$$

$AgNO_3$ limits! $= 0.235\text{g AgCl}$