How much solute is in your solution?

Learning Outcomes

1. Calculate grams required to make a solution of a desired molarity
2. Become proficient at making a solution using a volumetric flask
3. Use mass of a solid along with volume data to determine the molarity of the solution
4. Dilute a stock solution to a desired molarity
5. Titrate a solution to determine the molarity of the solution
6. Use the molarity of a known solution along with volume data and reaction stoichiometry to determine the molarity of a second solution

Introduction

The preparation of aqueous solutions is a critical skill needed for not only chemistry experiments but biology and biochemistry procedures as well. Solutions can be made by dissolving a solid in a liquid solvent or mixing two liquids. Creating a solution of the correct concentration is very important for the success of an experiment. Molarity (*M*) is the most common unit of concentration for solutions. Molarity is defined as moles of *solute* dissolved per liter of *solution* produced.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

The moles of solute can be determined either from the mass using molar mass or the volume of a given concentration of a solution. Therefore, this implies that a solution can be prepared two ways – one by dissolving a solid in water or the second by diluting a concentrated solution.

Because molarity is defined in terms of the volume of solution, special volumetric glassware called a *volumetric flask*, must be used for very *accurate* work. This glassware is carefully calibrated to contain a specific volume: 1000.00 mL, 500.00 mL, 10.00 mL, etc. The volume that a liquid occupies changes with temperature. Therefore, volumetric glassware is calibrated at a particular temperature (often 20^{o}C), which is approximately room temperature. Volumetric glassware should not, therefore be used to measure hot or cold liquids.

<u>Preparation of a Solution from a Solid or Pure Liquid Solute</u>

For a solution with a known *molarity*, the moles of the solute in the solution are easily calculated using the equation shown above. For example, 150 mL of a 1.50 M NaCl solution has 0.300 moles of NaCl, also referred to as the solute.

$$\text{Moles of NaCl} = 1.50\text{ M NaCl} = \frac{\text{x moles of solute}}{0.150\text{ L solution}} = 0.225\text{ moles of NaCl} \qquad \text{(Equation 1)}$$

Since moles is not a measurable quantity, the molar mass must be used to solve for the mass.

At this point the mass of the solute and volume of the solution is known. To prepare the solution a volumetric flask that measures the desired volume must be used. Since molarity is defined by liters of the total solution, the solid or liquid must be added to the flask first. Then water is added to fill the flask to the desired volume, as indicated by the etched line on the neck of the flask.

Preparation of a Solution by Dilution

Once a solution is carefully prepared using a volumetric flask and an analytical balance to accurately determine the mass of the solute, the molarity of the solution can be used to make dilute solutions. Very dilute solutions would be difficult to prepare if the mass is very low, even analytical balances have a limit!

The thought process behind the preparation of a dilute solution, is similar to the process described in the previous section. Ultimately, the number of moles of the solute in the dilute solution must be determined. Once this is found, the volume of the stock, or concentrated, solution that contains that quantity of the solute can be determined.

For example, one wishes to prepare 150.0 mL 0.0100 M KOH solution. As show in the calculations below this would require 0.0842 g of KOH. Although, an analytical balance could measure this quantity it me be difficult to get this precise amount since KOH is a pellet.

$$150 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) = 0.150 \text{ L solution} \qquad \text{(Equation 2)}$$

$$0.150 \text{ L solution} \left(\frac{0.0100 \text{ mol KOH}}{1 \text{ L solution}}\right) = 0.00150 \text{ mol KOH} \qquad \text{(Equation 3)}$$

$$0.00150 \text{ mol KOH} \left(\frac{56.11 \text{ g}}{1 \text{ mol KOH}}\right) = 0.0842 \text{ g KOH} \qquad \text{(Equation 4)}$$

Therefore, it may be more practical to prepare 500.0 mL of a 0.100M solution of KOH and then prepare the 0.01M solution. To achieve this, the 0.1M solution would be prepared using the method described in the previous section. Therefore, according to the calculations below 2.81 of KOH should be measured out (a much more reasonable quantity to measure out for this type of solid), and then transferred to the volumetric flask. The flask can then be filled with water to the etched line on the neck of the flask.

$$500 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) = 0.500 \text{ L solution} \qquad \text{(Equation 5)}$$

$$0.500 \text{ L solution} \left(\frac{0.100 \text{ mol KOH}}{1 \text{ L solution}}\right)\left(\frac{56.11 \text{ g}}{1 \text{ mol KOH}}\right) = 2.81 \text{ g} \qquad \text{(Equation 6)}$$

Now that the 0.100 M solution has been prepared, the volume of this concentrated solution that is needed to prepare the desired 0.0100 M solution must be determined. The ultimate question, as we have asked before, is how many moles of KOH are needed in the desired solution. Based, on the calculations performed in Equation 3, 0.01500 mols is required to prepare the dilute solution. Therefore, what volume of the concentrated solution contained 0.00150 mol of KOH? The calculations for how to solve this are shown below. As a result, 15.0 mL of the concentrated solution are required to supply 0.00150 mol of KOH to dilute solution.

$$0.00150 \text{ mol KOH} \left(\frac{1 \text{ L solution}}{0.100 \text{ mol KOH}}\right) = 0.0150 \text{ L or } 15.0 \text{ mL} \qquad \text{(Equation 7)}$$

If you have following the above calculations, you should have notice that the number of moles in required volume of the stock solution does not change when it is diluted with water. Therefore, the following equation can be used:

$$(\text{Molarity}_{\text{stock.soln.}})(V_{\text{stock soln.}}) = (\text{Molarity}_{\text{dilute soln.}})(V_{\text{dilute soln.}}) \qquad \text{(Equation 8)}$$

If you write out the complete units for molarity in Equation 8, you will find that this equation is simply stating that the moles of the solute from the stock solution is equal to the moles of the solute in the dilute solution. This is shown in Equation 9, where the volume on each side of the equation will cancel out.

$$\left(\frac{moles_{stock\ solution}}{Volume_{stock\ solution}}\right) V_{stock\ solution} = \left(\frac{moles_{dilute\ solution}}{Volume_{dilute\ solution}}\right) V_{dilute\ solution} \quad \text{(Equation 9)}$$

Example Problems Related to Solution Preparation

In this experiment you will be asked to prepare a NaOH solution of a given molarity starting with solid sodium hydroxide. The following example calculation shows how to determine the grams of a solute needed to prepare a solution of a particular concentration.

Sample calculation to determine grams of NaOH needed.
If 100.0 mL of a 0.075 M aqueous solution of NaOH is needed, how many grams of NaOH are required to dissolve in DI water?

a) Determine the moles of sodium hydroxide (NaOH) needed using molarity and volume:

Molarity of NaOH $\times Lsolution = 0.075M \times 0.1000L = 0.0075\, molNaOH$

b) Determine the grams of NaOH needed using molar mass:

$$\text{grams of NaOH} = 0.0075\ mol_{NaOH} \times 39.997\ g/mol = 0.30\,g\ NaOH$$

You will also make a second solution by diluting the first solution.

Sample calculation to create a dilute solution
What volume of the 0.075 M NaOH solution is needed to prepare 100 mL of a 0.040 M NaOH solution?

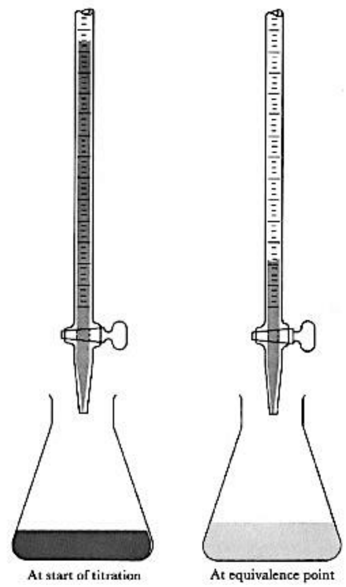
Dilution Formula:

(0.075 M)V = (0.040 M)(100 mL)

V = 53 mL

How to Determine the Concentration Experimentally

Once a solution is prepared, the concentration of the solution can be determined experimentally. By measuring the amount of product formed from a reaction, the amount of starting reagents can be determined. When a stoichiometric equivalent moles of acid and base react, they are said to be at the *equivalence point*. Since the acid and base will have equal moles at the equivalence point, the molarity and volume of the acid and base used can be used to calculate an unknown. The process of reaching the equivalence point with acids and bases is called *neutralization*.



In order to use a neutralization reaction to determine the concentration of a solution, a ***titration*** can be performed. For a titration, an acid of known concentration is slowly added to a known volume of base with unknown concentration. One method used to determine the equivalence point or neutralization point is the addition of an indicator to the reaction of the acid and base. An ***indicator*** is a compound that will react with one of the titration species and for which a color change is observed very near the equivalence point. The indicator used in this experiment is called phenolphthalein which is colorless in an acidic solution and pink when excess base is present. Once the volume of the added acid is determined, the concentration of the base solution is determined.

An example calculation for determining the concentration from a titration experiment is shown below.

Sample calculation

What is the concentration of a HCl solution if 12.78 mL are required to titrate 20 mL of a 0.050 M NaOH solution to the equivalence point?

a) Write the balanced chemical equation.

$$NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O$$

b) Determine the moles of sodium hydroxide (NaOH) in the flask using volume and moles:

Moles of NaOH = $molarity\ NaOH \times Lvolume = 0.050\,M \times 0.020\,L = 0.0010\ mol_{NaOH}$

c) Using the stoichiometric relationship in the balanced chemical reaction, determine the moles of base required to reach the equivalence point:

$$mol_{HCl} = mol_{NaOH} \qquad 0.0010\ mol_{NaOH} \times \frac{1\,mol_{HCl}}{1\,mol_{NaOH}} = 0.0010\ mol_{HCl}$$

c) Determine the concentration of HCl by dividing the number of moles by volume of HCl solution added:

molarity of HCl = $\dfrac{0.0010\ molHCl}{0.01278\ L_{HCl\ solution}} = 0.078\,M\ HCl$

Procedure

<u>Preparation of the NaOH Solution:</u>

Calculate the approximate amount of sodium hydroxide needed to make 250 mL of a NaOH solution of a given molarity (provided by your instructor). Have your instructor or TA check your calculations before proceeding.

Obtain a 250 mL volumetric flask from the supply shelf. Also, obtain a 250 mL bottle and place a small piece of labeling tape on the side of it. Include on the label: the concentration and name (written out, not abbreviated) of chemical in solution, your name, and the date of preparation. (This type of information should be included on every solution that you prepare in a lab).

In a weigh-boat, obtain a mass of NaOH pellets that is <u>close</u> to the mass you calculated for the NaOH solution. The mass does not need to be exactly what was calculated, but you should be within a few tenths of a gram. Record the exact mass.

Hold the weigh-boat up to the edge of the volumetric flask that you labeled and carefully transfer the contents of the weigh-boat into the flask. Fill the labeled 250-mL volumetric flask about one-half of the way with DI water. You may use a “DI”-wash bottle to wash the contents of the weigh boat into the bottle.

Cap the volumetric flask and mix the contents with gentle swirling until all of the pellets have dissolved. (This mixing process is exothermic so the container may get warm). Next fill the bottle with DI water to its marked line that indicates 250 mL, recap, and again gently swirl until mixing is complete.

Pour the contents into the labeled bottle and wash out the volumetric flask.

<u>Dilution of the NaOH Solution</u>

Next, a solution that is one-fifth (1/5) the molarity of the assigned molarity in the previous section will be prepared. To do so, begin by determining what the concentration of the dilute solution will be. Using the dilution formula in Equation 8, determine how much of the concentrated NaOH solution (that was prepared in the previous step) is required to prepare 250mL the new dilute solution. Have your instructor or TA check your calculations before proceeding.

Using a graduate cylinder, measure the appropriate volume of the original NaOH solution to create the diluted solution. Pour the desired volume of the NaOH solution into the cleaned 250 mL volumetric flask and dilute with DI water to the 250 mL fill line. Swirl to mix efficiently.

Pour the contents into a second **<u>labeled</u>** 250 mL bottle and wash out the volumetric flask.

<u>Titration of the NaOH Solution</u>

Determine the appropriate HCl solution to use to titrate your original NaOH solution by calculating the approximate volume of each available HCl solution needed to have equal moles with the moles of NaOH in 10 mL of the solution. Have your instructor or TA check your calculations before proceeding. The correct HCl solution should require 5 to 20 mL to titrate.

Obtain a buret, ring stand, and buret clamp. Your instructor or TA will demonstrate how to put the buret stand together.

With the stopcock of the buret closed, rinse the buret twice with 5 mL of the correct HCl solution. The HCl solution can be poured into the top of the buret using the plastic funnel in your drawer. This action will ensure that no residual chemicals from a previous class are in the buret to affect the purity of the HCl solution. Allow the rinse solution to drain through the stopcock into a waste beaker. Complete this step a second time and set the waste beaker aside. With the stopcock closed, fill the buret with the a sufficient volume of the HCl solution to complete two titrations. (Use your calculations as a guide and do not be wasteful!)

Open the stopcock, letting some of the HCl solution flow out in your waste beaker, so as to fill the tip of the buret, eliminating air in the tip; then close the stopcock. The solution level need not be at 0.00 mL but at some level where you can read the volume with your eyes even with the bottom of the meniscus. Remove the funnel and record the volume reading of the buret **to the nearest 0.02 mL**. This volume reading will be the "initial volume" for your titration.

Add 10 mL of your original NaOH solution to a clean 100 mL Erlenmeyer flask using a volumetric pipet. Add 2-3 drops of a 1% phenolphthalein indicator solution to the flask.

Begin the titration by adding the HCl solution slowly, drop by drop, SWIRLING the flask continually. Continue titrating until the solution turns from pink to colorless. The end of the titration is indicated when the pink color is completely gone. Record the final volume on the buret to the nearest 0.02 mL.

Repeat the titration again to verify the results, using another 10 mL of the NaOH solution. Record the volume of the HCl solution required.

Next, titrate the diluted solution. Determine the appropriate HCl solution to use. **It will not be the same HCl solution as used previously!** Check your selection with your instructor or TA.

Repeat the titration steps again with the diluted NaOH solution. Once completed you will have performed 4 titrations: 2 with the original NaOH solution and 2 with the diluted NaOH solution. Each titration should start with 10 mL of the NaOH solution.

Use the data from the titrations to determine the molarity of the original NaOH solution and the diluted NaOH solution.