1.0 Introduction

Acid base titrations are often used to determine the concentrations of unknown acids or bases. In order to accomplish this, a standard must be available. A standard is a solution of known concentration. Primary standards are stable substances of high purity that can be used to prepare standards directly. The primary standard can then be used to standardize another solution which is not a primary standard. This is necessary because some substances are not stable or are not available in sufficient purity to directly prepare a standard solution from them. In this case, a solution of approximate concentration is prepared and is then standardized with another standard. This is the case with the unknown acids prepared for this lab. Solutions of acetic acid (vinegar) or hydrochloric acid are not primary standards so the solutions are made to an approximate concentration and must then be standardized. The base used in this experiment, sodium hydroxide, is not a primary standard either but it has a known concentration of 0.100 M which was previously determined.

The concentration unit for solutions often used by chemists is molarity or moles of solute per liter of solution. This is a direct measure of the number of solute particles dissolved in a unit volume of the solution. The abbreviation for molarity is M which for some unit or dimensional analysis problems should be converted to : moles substance / L soln. Since we are using a known volume of the acid and know the concentration of the standard NaOH solution, we can measure the volume of NaOH used to complete the titration and calculate the molarity of the unknown acid using molarity and volume relationships. To understand more details consider the following reaction.

When acids and bases react, they neutralize each other to form a salt and water.

$$HX + BOH \longrightarrow BX + H_2O$$

When an equal amount of acid and base have been mixed together and the solution is completely neutralized we are at the equivalence point. Since acids and bases are usually colorless, how can we tell when equal amounts of the acid and base are present? In practice, we use indicators which are often acidic substances which undergo color changes as they lose or regain their acidic protons (hydrogens). As long as the solution is acidic the indicator is color one but as soon as the solution begins to turn basic color two appears. The point at which an indicator changes color is its endpoint. Indicators are chosen so that their endpoints align with the equivalence points of acid base titrations. Thus, a change in indicator color indicates the equivalence or neutralization point of the titration. Notice that you must stop titrating at the first persistent sign of the indicator color change or else you may over-titrate past the true equivalence point of the titration!

By equal amounts of acids and bases, we really mean equal numbers of hydrogen ions and hydroxide ions and not equal masses of the substances. This is complicated by the fact that some acids have multiple protons and some bases have multiple hydroxides i.e. H_2SO_4 and $Ca(OH)_2$. For our purposes we will limit our focus to acids and bases with only one reactive unit. If the number of acid units (moles) equals the number of base units (moles), then the solution will be completely neutralized. Analysis of the units shows that M = mole/L, therefore moles base/L X V in L cancels the liter units leaving only moles base used. Since we know moles acid = moles base, we also know the moles of acid at this point. Using this value and the known volume of acid enables us to calculate the molarity of the unknown acid.

Notice that the assumption that moles of acid = moles of base is NOT true for any amounts of acid and base used, but ONLY at the equivalence point of the titration.

In this experiment you will use either unknown hydrochloric acid, HCl, or unknown acetic acid, $HC_2H_3O_2$, and titrate these with a standard solution of sodium hydroxide, NaOH. The indicator, phenolphthalein, will turn from colorless in the acid to pink at the equivalence point. At that point the titration is complete and you can calculate the M for your acid.

2.0 Procedure

A. Clean and prepare your buret as suggested by your instructor. Normally the last step in the cleaning process will be to rinse the buret with small quantities of deionized water. DO NOT forget the tip ! Follow this by several rinses with small amounts (5mL) of the NaOH. Now fill your buret with the standard NaOH solution to near the 0.0 mL mark. Run some of the NaOH solution through the buret tip and make sure the tip is full at the beginning of the titration. Place the buret securely in a buret clamp.

B. Pipet 10.0 mL of the unknown acid (option 1 or 2) into a 250 mL Erlenmeyer flask, add about 50 mL of deionized water (exact volume not important) and add 2-3 drops of the phenolphthalein indicator.

C. Read and record the initial volume of the NaOH in the buret. This volume does not have to be 0.0 mL but might be any value near the top of the buret. Make sure to read with your eye level to the buret. This may require lowering the buret or standing on a stool for a short person.

D. Begin adding the NaOH to the flask slowly with swirling. You DID remember the indicator didn't you ? Watch for faint "flashes" of pink which will grow more persistent as the titration nears the end point. It might help to place a white paper towel under your flask so the black desktop does not interfere with seeing the pink color. Continue to titrate until the first sign of pink color that persists for 30 seconds of swirling develops. You don't want to go too far - a persistent faint pink color is perfect.

E. Read and record the final volume on your buret. It might help to hold a white towel or card behind the buret to block out background interference.

F. Using clean (wet with deionized water is fine) flasks, repeat steps B-E two more times. Refill your buret with NaOH solution as needed. Notice that each titration should take the same amount of base since you are always using 10.0 mL of the acid. You can speed up the titrations by adding NaOH rapidly until you are within a couple of mL of the end point. Now you will only have to add 20 to 40 drops to hit the end point so it will be much easier.

G. When you have two titrations which agree within 0.2 mL of each other you can stop. Otherwise, try another titration. If you are having difficulty, a consultation with your instructor would be wise.

H. When your results are acceptable, clean your glassware and return them to the storage area. DO NOT RETURN GLASSWARE WITH ACIDS AND BASES STILL IN THEM.

3.0 CALCULATIONS

Use your titration data, known molarity of the standard NaOH (write this on the report sheet), volume of NaOH used and the 10.0 mL volume of the acid , to calculate the molarity of the acid.

4.0 Report Sheet

UNKNOWN I D _____ **Molarity of Standard NaOH _____**

Sample Number	1	2	3	4
Initial buret reading (mL)				
Final buret reading (mL)				
Volume of NaOH used (mL)				
Molarity of Unknown Acid				

QUESTIONS:

1. Why must you use a pipet bulb to draw liquids into the pipet ?

2. Give the calculation setup that shows that, if 10.0 mL of HCl(aq) requires 16.2 mL of 0.403 M NaOH in a titration, the molarity of the HCl is 0.653 M.

3. Calculate the molarity of an acetic acid solution, $HC_2H_3O_2(aq)$, if a 50.0 mL sample requires 35.8 mL of 0.150 M NaOH in a titration.