TITRATION OF ACIDS AND BASES

Reminder – Goggles must be worn at all times in the lab!

PRE-LAB DISCUSSION:

In the chemistry laboratory, it is sometimes necessary to experimentally determine the concentration of an acid solution or a base solution. A procedure for making this kind of determination is called an ACID-BASE TITRATION. In this procedure, a solution of known concentration, called a STANDARD solution is used to neutralize a precisely measured volume of the solution of unknown concentration to which one or two drops of an indicator have been added. Since the solution of unknown concentration is acidic, a standard base solution is added to the acid solution until it is neutralized.

When carrying out an acid-base titration, you must be able to recognize when to stop adding the standard solution. That is, you must be able to recognize when neutralization has occurred. This is the purpose of the INDICATOR. A sudden color change due to the indicator signals that neutralization has occurred. At this point, the number of HYDRONIUM ions from the acid is equal to the number of HYDROXIDE ions from the base. The point at which this occurs is called the END-POINT of the titration. When the endpoint is reached, the volume of the standard solution is carefully determined. Then the measured volumes of the two solutions and the known concentration of the standard solution can be used to calculate the concentration of the other solution.

At the end-point of the experiment, the volume of the acid times the molarity of its hydrogen ions will equal the volume of the base times the molarity of its hydroxide ions. Calculations will be based on the following formula:

$$(\text{volume of acid})(\text{molarity of } H^+) = (\text{volume of base})(\text{molarity of } OH^-)$$

or, more simply…

$$V_a M_{H+} = V_b M_{OH-}$$

Knowing three of these quantities we can calculate the fourth -- in this case the molarity of the acid's hydrogen ions. Since the acids and the base in this experiment have one hydrogen ion and one hydroxide ion, respectively, the molarity of hydrogen ion is the same as the molarity of the acid, and the molarity of hydroxide ion is the same as the molarity of the base.

PURPOSE:

To learn the experimental technique of titration and using this technique, to determine the molarity of three acid solutions of unknown concentration.

PROCEDURE:

<u>PART I: Titration of HCl "A" solution</u>.

1. The burets have been set up for you at your work station and contain 0.20 M NaOH. If you should need more 0.20 M NaOH, the instructor will refill it for you. When reading the burets, be sure you always read the <u>bottom of the meniscus</u>!
2. For this lab, you will need to use distilled water. Do not use tap water! If your distilled water bottle needs to be re-filled, please re-fill it at the side counter. If you do not know how to dispense the distilled water, ask for help from your instructor.
3. Take your erlenmeyer flask from your drawer and wash it well. Once the flask is clean, rinse down the sides with distilled water. There is no need to dry the flask.
4. Take the flask to the side table and get EXACTLY 1.00 ml of HCl A solution. Be sure to read the bottom of the meniscus! Record this volume of acid in the Data section.
5. Go back to your station and rinse down the sides of your flask with several milliliters of distilled water. This is to ensure that all the hydronium ions are down in solution where they can be titrated.
6. Add 3 drops of bromthymol blue indicator to the contents of your Erlenmeyer flask. The solution in the flask should turn yellow.

7. Now you are ready to titrate. Looking at the buret at your station, on scratch paper, record the initial reading of the 0.20 M NaOH solution. Remember that volumes can be read <u>2 places past</u> the decimal point. Place your flask containing the acid below the buret. Now add the NaOH a little at a time, carefully opening the pinchclamp on the buret. Swirl the flask with a rotating motion after each addition of the NaOH. When near the end-point (neutralization), add one drop at a time. The end-point is reached when the entire solution is blue, and the color remains, even after swirling. At the end-point, read the final volume of base in the buret and record the volume of base used (SUBTRACT) in the Data section. You will also want to record the molarity of the base which is 0.20 Molar for all titrations in this experiment.
8. After the first titration, wash out the flask again, rinse down the sides with distilled water, and begin another trial. Continue to run trials until you get <u>THREE trials that agree within 0.20 ml of base used</u>. I show room in the sample data table for 6 trials – Hopefully it won't take that many (but it might take more). Circle the three accepted trials in your data table when you are done.

<u>PART II: Titration of HCl "B" solution</u>.
1. Follow the same procedure as in Part I, using HCl "B" solution instead of HCl "A" solution. As in Part I, continue doing trials until you get three trials that agree within 0.20 ml of base used. Circle the three accepted trials in your data table when you are done.

<u>PART III: Titration of Vinegar</u>.
1. Follow the same procedure as in Part I and II EXCEPT you will be using phenolphthalein as your indicator. Look for a color change from clear to very light pink to indicate the end point of the titration. Circle the three accepted trials in your data table when you are done.

RESULTS:

<u>Data and Observations</u>
You will need ***three*** identical tables like the one below in order to record all of your data!

Trial	1	2	3	4	5	6
Vol. Acid						
Vol. Base						
Molarity of Base						

<u>Calculations</u>

1. Showing ALL of your work, calculate the AVERAGE molarity of the three acceptable trials for HCl A solution. Repeat the calculations for HCl B, and vinegar, recording your final results in a table similar to the one at the right.

Solution	**Average Molarity**
HCl "A"	
HCl "B"	
Vinegar	

2. By law, vinegar must be LESS than 5% acetic acid. White vinegar tends to be on the high end. Apple cider vinegar tends to be less, and therefore, people prefer its taste. Attempt to calculate the % of acetic acid in your vinegar samples. You must show all of your work, including units!
   a) Calculate the <u>average</u> **molarity** of the three accepted trials of your **vinegar** titrations.
   b) Calculate the molar mass of acetic acid ($HC_2H_3O_2$).
   c) Calculate the grams of acetic acid per liter by multiplying the molarity by the molar mass.
   d) Convert your answer to a percentage. <u>Hint:</u> Right now you should have grams of acetic acid per 1000 ml. Move the decimal point so that you have grams per 100 ml.