Lab 19: Titration of Acids and Bases

**Background**: Titration is a method used to determine an unknown concentration of either an acid or a base using the known concentration of a base or acid, respectively. Adding enough titrant (solution in buret) to make the solution in the beaker change color along with the titration equation will be necessary to find the unknown concentration.

**Materials:**

Labeled Diagram of Experimental Setup

* Beaker
* Graduated cylinder
* Ring Stand
* 2 burets
* Buret clamp
* Phenolphthalein
* NaOH (0.2M)
* HCl (unknown molarity)
* H2SO4 (0.2M)
* Safety glasses
* Apron

**Procedure:**

**PART 1**

1. The buret contains HCl of an unknown Molarity (concentration.)
2. Obtain 10 ml of 0.2 M NaOH in a beaker. Put a few drops of Phenolphthalein in the beaker and swirl it. Your solution should turn a nice pink.
3. Record the initial volume of HCl in the buret in table 1. Remember the numbers are reversed.
4. Carefully turn the stopcock knob and **slowly** titrate the HCl into the beaker with NaOH. Every so often, you will need to swirl it to thoroughly mix the liquids.

\*\*When the mixture is getting close to changing colors, add even less acid at a time and keep swirling the beaker.

1. When the mixture in the beaker is colorless, stop the titrating. Record the final volume of HCl in the buret in table 1.
2. Thoroughly rinse out and dry your beaker, and repeat steps 2-5 for trials 2 and 3.

**PART 2**

1. Have your teacher pour 0.2 M H2SO4 into the buret on the right.
2. Obtain 10 ml of 0.2 M NaOH in a beaker. Put a few drops of Phenolphthalein in the beaker and swirl it. Your solution should turn a nice pink.
3. Record the initial volume of H2SO4 in the buret in table 2.
4. Slowly titrate the H2SO4 into the beaker with NaOH. You will need to swirl it to thoroughly mix the liquids. When the mixture in the beaker is colorless, stop the titrating. Record the final volume of H2SO4 in the buret in table 2.
5. Thoroughly rinse out and dry your beaker, and repeat steps 2-4 for trials 2 and 3.

**Data:**

**Table 1:**

|  |  |  |  |
| --- | --- | --- | --- |
|  | Trial 1 | Trial 2 | Trial 3 |
| Initial volume of HCl in buret. |  |  |  |
| Final volume of HCl in buret. |  |  |  |
| Volume of HCl used |  |  |  |

**Table 2:**

|  |  |  |  |
| --- | --- | --- | --- |
|  | Trial 1 | Trial 2 | Trial 3 |
| Initial volume of H2SO4 in buret. |  |  |  |
| Final volume of H2SO4 in buret. |  |  |  |
| Volume of H2SO4 used |  |  |  |

**Data Processing:**

1. To determine the Molarity of the HCl in the experiment, use this formula:   
   Molarity of the Acid \* Volume of the Acid = Molarity of the Base \* Volume of the Base   
    **Or simply MA VA = MB  VB**   
   1. Calculate the Molarity of HCl in trial 1: **SHOW YOUR WORK (formula, set up, answer)**
   2. Calculate the Molarity of HCl in trial 2: **SHOW YOUR WORK (formula, set up, answer)**
   3. Calculate the Molarity of HCl in trial 3: **SHOW YOUR WORK (formula, set up, answer)**
2. If the accepted value of the molarity of HCl is 0.1 M, calculate the percent error in the experiment. Use an average of your calculated values in question #1 for your measured value. **Show the formula, set up and answer.**
3. Write a balanced chemical equation for the neutralization reaction in part 1.
4. Write a balanced chemical equation for the neutralization reaction in part 2.
5. In the neutralization reaction in part 2, equal concentrations (molarities) of NaOH and H2SO4 were used.
   1. Which chemical did you need more of?
   2. Why do you think you needed more of one of the chemicals if they had equal molarities?
6. What are 2 specific sources of error in this experiment?