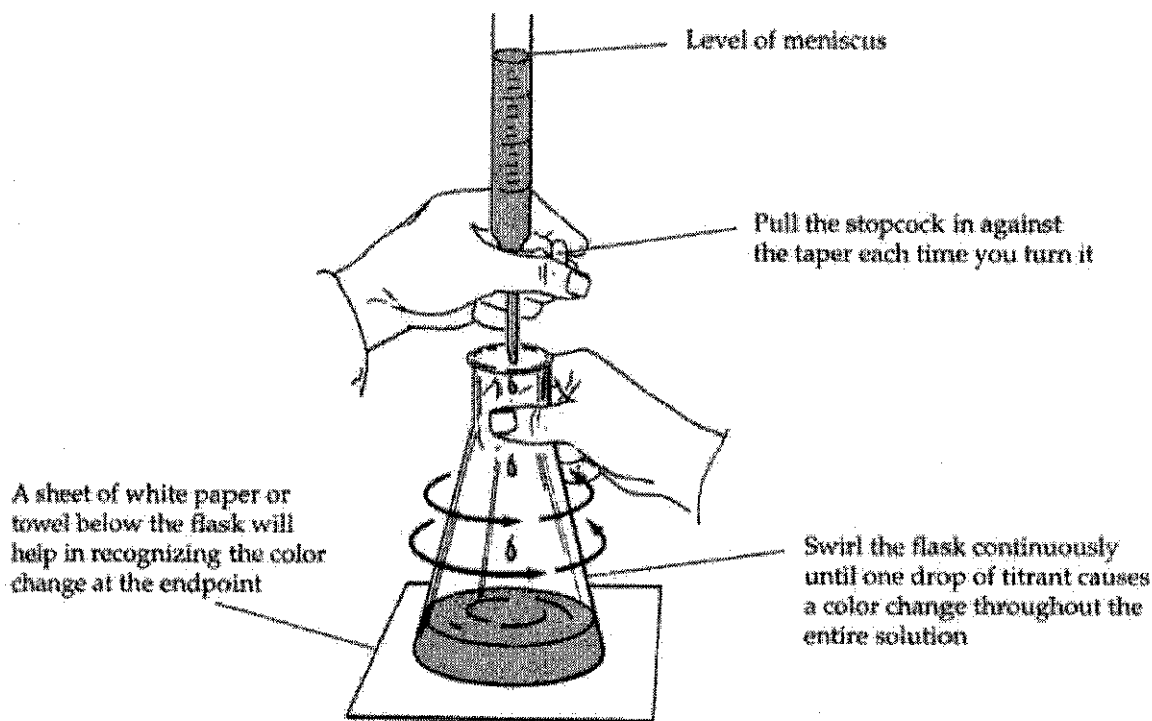


AP Chemistry Lab #2 – Inquiry Acid/Base Titrations

Introduction: One of the most common and familiar reactions in chemistry is the reaction of an *acid* with a *base*. This reaction is called *neutralization*, and the essential feature of this process in aqueous solution is the combination of hydronium ions (H_3O^+) with hydroxide ions (OH^-) to form water. In this experiment, this reaction will be used to determine accurately the concentration of a hydrochloric acid solution. The process of determining the concentration of a solution is called *standardization*. The technique of accurately measuring the volume of a solution required to react with another reagent is termed *titration*. The titration process is shown in Figure 20.2 below and is described by your instructor prior to experimentation.



▲ FIGURE 20.2 Titration procedure.

An *indicator solution* is used to determine exactly when an acid has neutralized a base or vice versa. An indicator changes colors when equivalent amounts of acid and base are present. The color change is termed the “*end point*” of the titration. Indicators change colors at different pH values. *Phenolphthalein*, the indicator used in this experiment, changes from colorless to pink at a pH of 9. In slightly more acidic solutions it is colorless, and in more alkaline (basic) solutions it is pink. In this experiment, sodium hydroxide will be titrated against an unknown solution of HCl and a measured amount of citric acid.

1. Preparation of a standard NaOH solution:

Mass out [*approximately*] enough solid NaOH to prepare 100.00 mL of a 0.150M NaOH solution. Make sure you record the exact mass of NaOH used and then show the set-up and calculations for the actual concentration of your standard solution based on the *actual* mass of NaOH you weigh out on the balance.

- NaOH is both caustic and hygroscopic. (*What does this mean?*)
- Due to the size of the solid NaOH pellets, obtaining an exact mass is not necessarily possible, but the exact mass used is important. (*Why is this true?*)

- Transferring of as much of the base as possible to your beaker is crucial...then from the beaker to the volumetric flask as well! (*Examine the procedure...where would this be especially important?*)
- Prepare or 'Standardize' your burette (follow the procedures outlined by your instructor).

Procedure:

- a) Mass the NaOH in the plastic weigh boat. (*What is the problem with chemicals directly touching the balance?*)
- b) Transfer the solid NaOH to a small beaker (*you are going to make 100 ml...which size beaker should you choose and how much water should be put inside this beaker?*)
- c) Dissolve the solid in the beaker and then transfer the entire contents to a volumetric flask and bring the total volume to the mark (Choose your beaker wisely and transferred all of the NaOH. *Explain the importance of this hint.*!)
- d) Seal your volumetric and mix. Introducing the miracle material called *parafilm*.
- e) 'Standardize' your burette with your NaOH solution. (Per instructor directions.)
- f) Set up the burette and prepare the needed glassware (standard beaker, waste beaker, white paper, rinsing squirt bottle).

2. Titration of the standard NaOH vs. a known concentration of HCl:

Titrate your standard vs. 10.00 mL of a known HCl solution. Report your findings based on this titration.

- This will allow you to practice your titration technique
- Sometimes a "back titration" is used where another burette containing acid is used to "back-up" past the endpoint...or you may add more acid (measured) if you pass the endpoint

Procedure:

- a) You should already have a general idea of the volume of NaOH needed to neutralize the given volume of acid...if you do not have any idea...determine it now before you start!
- b) Place exactly 10.00 mL of acid into a 250mL Erlenmeyer flask along with **about** 10-15 mL of water. (*Why is this volume of water unimportant?*)
- c) Add one to two drops of the phenolphthalein indicator to the flask and swirl to mix. (*Why not use more...like say a few mL or just pour some in?*)
- d) Following proper titration techniques (aka. instructor directions), begin adding the NaOH solution to the flask while swirling.
- e) Decrease the amount of each addition of base as you near the endpoint. (*What is the endpoint and how is it related to the equivalence point?*)
- f) Occasionally wash the insides of the flask down with water from the squirt bottle. (*Explain this step.*)
- g) You have successfully reached the endpoint if you are able to add either a single drop or "split" a drop into your flask and the pink color persists for about 30 seconds before fading to close to clear. If you don't know how to 'split a drop'...ask your instructor!
- h) If your solution is too pink...you have gone too far (consult the "back titration" procedure or start over).
- i) Rinse the contents of the flask down the sink with water. Repeat the titration if needed.

3. Determination of the molar mass of an "unknown" triprotic acid:

Determine the molar mass of citric acid by titrating a 0.20 g sample vs. your standard. The exact mass is very important...but having exactly 0.20 g is not! Follow the same procedure outlined above!

Clean-up: Triple rinse all glassware and return the burette to the holder (upside down to drain).

Acid-Base Titration Lab

Pre-Lab

1. Write out the neutralization reaction for sodium hydroxide reacting with hydrochloric acid.
 - a. Complete ionic equation
 - b. Net ionic equation

2. Calculate how many grams of solid NaOH will be necessary to prepare 100.00 mL of a 0.150M NaOH solution.

3. What volume of 0.150M NaOH is necessary to titrate 10.00ML of 0.1M HCl?

In/Post Lab

4. When completing the titration, how many mL of 0.150M NaOH did it take to neutralize the 0.1M HCl solution?

5. What materials are necessary to complete this experiment? List them.

6. Include a picture of your citric acid/NaOH titration.

7. Write out the procedure you used to titrate citric acid. Make sure the following terms are (correctly) used in your procedure: volumetric analysis, titration, equivalence point, indicator, endpoint, titrant, and analyte.

8. Calculate the molar mass of a 0.2 gram sample of citric acid based on your titration data.

9. Calculate the percent error for the molar mass of citric acid.

10. Write a conclusion based on your data and error analysis.