Experiment 11 - Acid-Base Titration

Introduction

A **titration** is an experimental technique for determining the **molarity** of a substance (the analyte) in solution by reacting it with another substance (the titrant) for which you can keep track of the amount, or volume, of titrant reacted. Recall that molarity of a solution is defined as moles of solute per liter of solution, so a 1 M ("one molar") solution has 1 mole of solute in 1 liter of solution. In this case the analyte will be acetic acid, and the titrant will be NaOH (for which you will be provided a solution of known molarity, i.e. a standardized solution).

To perform a titration, a carefully measured volume of the analyte is added to an Erlenmeyer flask. An **indicator** is added that will signal the **endpoint** of the titration by a visible color change. Then the titrant is added slowly to the flask using a buret. When the indicator changes color, exactly all of the analyte has been used up. At this point, you assume that you have added exactly the right number of moles of the titrant to completely react with the analyte (the mole ratio used exactly matches the ratio in the balanced equation). Since you know the volume of analyte you started with, and the volume of titrant added (using the buret) to reach the endpoint, you can easily calculate the analyte concentration of the unknown solution using stoichiometry.

Titrations are often performed to measure concentration of an acid (or a base) in solution. Acid-base titrations are convenient because there are many appropriate indicators. In this lab, we will use **phenolphthalein** as our indicator. Phenolphthalein is clear (colorless) in acidic solution, but pink in basic solution. In each titration, we will put the acidic solution containing the analyte in Erlenmeyer flask, and titrate in the basic solution containing the titrant from the buret. The solution in the flask will start clear and turn pink when the endpoint is reached.

The ultimate goal of this lab is to measure the concentration of acetic acid in vinegar.

Safety Precautions:

- Wear your safety goggles.
- If any acid or base solution splashes on you, rinse it off immediately.

Waste Disposal:

- Waste from this experiment may be safely discarded down the drain using plenty of running water.

Procedure

Mass Percent of Acetic Acid in Vinegars

You will determine the precise acetic acid content in a specific brand and type of vinegar by titration with provided standardized NaOH. You will do at least 3 titrations, repeating the titrations until subsequent titrant volume results do not differ by more than 1.5%.

1. Collect a small amount of vinegar in a beaker or small flask (each person will need about 15 mL, enough for several titrations). You will use 2 mL volumetric pipets to deliver vinegar to Erlenmeyer flasks. Condition your volumetric pipet with ~ 2 mL of vinegar. Make sure to use the same brand of vinegar for the whole experiment. Ask for a tutorial if you have never used a volumetric pipet and one of us will be glad to demonstrate.
2. Condition a large buret with the standardized NaOH solution.

3. Weigh a clean, dry Erlenmeyer flask (100 mL is fine) on an analytical balance to a precision of at least ± 0.001 g and record the mass in your notebook.
4. Pipet about 2.00 mL of vinegar to the flask with a volumetric pipet. Weigh the flask again and record the mass of the flask with the vinegar. Subtract to determine the mass of vinegar used.
5. Add a few drops of phenolphthalein indicator into the flask with the vinegar. If the vinegar has a dark color, add more purified water to lighten it.
6. **Use the provided standardized NaOH solution** (note the NaOH molarity for your calculations) to titrate the vinegar. Be sure to record the initial buret reading, as well as the buret reading at the endpoint, to two decimal places.
7. After the first titration, dump the contents of the flask down the sink, rinse the flask well with deionized water, shake it dry, and carefully dry off the outside of the flask with a paper towel.
8. Repeat the procedure from step 3-7 at least twice. (Weigh the flask, add vinegar, re-weigh the flask, and titrate. The flask does not have to be absolutely dry, but it should be dry on the outside. You will need to weigh it before each trial, because it will contain a slightly different amount of water each time. Alternatively use several dry flasks.)
9. Calculate the concentration of the vinegar and check the results with your instructor. If necessary, repeat the titrations until your results are close enough together.
10. Once your instructor approves your results, clean up. The waste can go down the sink.

Calculations for Part

1. Calculate the concentration of the vinegar separately for each trial, as you did for Part 2, using the precise concentration you found for the NaOH.
1. Find the % difference between your least similar trials (of 3 trials total). If it is not less than 1.5%, do more trials until you have 3 within 1.5% of each other.
2. Find the average concentration (in molarity) and the average deviation of the vinegar.
3. Find the density of the vinegar from each trial. Calculate the average density and average deviation.
4. Using the average concentration and the average density, find the mass % of acetic acid in the vinegar.
5. Record the density, molarity and mass % with their error ranges (as described in the lab notebook guidelines).

Name: Section:

Experiment 11 Pre-Lab Sheet:

1. Suppose you are titrating an unknown amount of acetic acid, CH_3COOH, in a flask, and it takes 23.02 mL of 0.1965 M NaOH to reach the equivalence point. How many moles of acetic acid were originally in the flask? If there were originally 10.00 mL of the unknown acetic acid solution, what was the original concentration of acetic acid?

2. In the previous exercise, the 0.1965 M NaOH solution is called the "standardized base". What does "standardized base" mean? In general, what is the definition of a standardized solution? Most importantly, how would you go about standardizing an NaOH solution of unknown molarity?

3. How and why must you condition your burets and pipets?

4. Find the concentration of a solution of sulfuric acid if it requires 30.45 mL of 0.1048 M NaOH to titrate 44.81 mL of the sulfuric acid solution. (Hint: write the balanced equation! Why can't you use the dilution rule?)