Activity 27: Mass Titrations - Measuring Molar Concentrations[1]

Goals

- **Measure** the molar concentration of acids using mass titrations.
- **Compare** the accuracy of mass titrations with that of volumetric titrations.

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. What is the difference between mass and weight?
2. What is a titration?
3. What is a *dilute aqueous solution*?
4. Why do chemists do more than one trial when doing quantitative analysis?
5. What type of chemical reaction occurs when sodium hydroxide is added to hydrochloric acid?
6. What does *%(w/w)* mean?

Concepts to Review

Definitions of Mass
Chemical Reactions
Solutions

Introduction

A *mass titration* (also called a *weight titration)* is a method of finding molar concentration by weighing solutions rather than measuring their volumes. A mass titration is often more accurate and precise than a volumetric titration because a balance is usually a more accurate and precise instrument than a pipet or buret. The result of a mass titration depends only on weighings determined directly from the balance and not on volumes determined from pipets and burets.

You can use the balance to determine the volumes of solutions by weighing. You know that the density of water is 1.00 g/mL. The densities of dilute aqueous solutions can be safely assumed to be equal to the density of water. This means that the mass of a solution in grams is numerically equal to its volume in mL.

In this lab you will carry out several mass titrations. You will weigh a certain amount of acid solution and add an indicator. Then you will determine the mass of NaOH solution needed to reach the end point. To do this you will start by weighing a transfer pipet full of NaOH solution. There is no need to weigh the empty pipet, as you will only need to know the *difference* in mass before and after the titration. Next, you will carefully add NaOH solution to the acid solution until the indicator just changes color. Finally, you will determine the mass of the NaOH solution needed for the titration by weighing the NaOH pipet again, and taking the difference in the initial and final weighings.

Notice that you do not have to count drops, nor do you have to hold the pipet at any particular angle, nor do you have to be careful to expel any air bubbles. In short, you will work with none of the disadvantages of a drop-counting, "digital" titration.

[1] Adapted from: Waterman, E. L. *Chemistry: Small-Scale Chemistry Laboratory Manual;* Addison-Wesley/Prentice-Hall, Inc.: Upper Saddle River, New Jersey, 2002; pp 87-92.

Safety

Wear safety glasses. Both sodium hydroxide and the acid solutions are *very* dangerous if you get a drop into your eye.
Act in accordance with the laboratory safety rules of Cabrillo College.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Small scale pipets of the following:

0.5 *M* NaOH	unknown HNO_3	unknown HCl
phenolphthalein	vinegar	
unknown H_2SO_4	lemon juice	

Equipment: Balance Plastic cups

Experimental Procedure

1. Weigh, to the nearest 0.01 g, approximately 0.5 g of acid solution as follows:
    a. Place an empty plastic cup on the balance pan. You can rinse it with deionized water from a wash bottle. Make sure the cup is dry on the outside, but it need not be dry on the inside. Then "tare" the balance so that it reads 0.00 g.
    b. Add acid solution from the transfer pipet until the balance gives a reading close to 0.5 g. Then record the mass of the acid solution to the nearest 0.01 g (i.e., record every digit you see on the readout).
2. Add 3 drops of phenolphthalein.
3. Place *two* full NaOH pipets in a second clean, dry plastic cup, and record their initial mass. (Use two pipets to make sure you have enough NaOH.)
4. Titrate the acid solution in the first plastic cup with NaOH from one or both of the pipets until *1 drop* turns the solution to a stable pink color. "Stable" means the pink color doesn't fade when you stir the titration solution, and it stays pink for at least 30 seconds.
5. Determine the mass of the NaOH solution used for the titration:
    a. Place *both* NaOH pipets back into the same clean, dry plastic cup used before.
    b. Weigh the cup and the pipets, and record their final mass.
    c. Subtract the final mass from the initial mass. This gives the net mass of NaOH solution delivered into the titration cup.
    d. Record the final mass as the initial mass for the next titration, if necessary.
6. Repeat steps 1-5 until you obtain consistent results. (The only way to tell if your result is reliable is to reproduce it.) Calculate the difference between your highest and lowest results for a given acid as a percentage of their average. This value should be less than 1%.
7. Carry out mass titrations to determine the molar concentrations of HCl, HNO_3, and H_2SO_4. Titrate each acid at least twice. Use the third column in the data table if necessary.
8. Next titrate the acetic acid in vinegar, and the citric acid in lemon juice. Again, titrate each solution at least twice.

Activity 27: Mass Titration

Name______________________________

Section_____________ **Date**_______

Experimental Data:

Record the precise molarity of the NaOH solution: ___________*M*.

Molarity of Laboratory Acids

	HCl			HNO_3			H_2SO_4		
Mass of acid (g)									
Initial mass of NaOH (g)									
Final Mass of NaOH (g)									
Mass of NaOH used (g)									
Molarity of acid (*M*)									
Average acid molarity	*M*			*M*			*M*		

Molarity of Acids in Household Products

	$HC_2H_3O_2$ (acetic acid in vinegar)			$H_3C_6H_5O_7$ (citric acid in lemon juice)		
Mass of acid (g)						
Mass of NaOH (initial) (g)						
Mass of NaOH (final) (g)						
Mass of NaOH used (g)						
Molarity of acid (*M*)						
Average acid molarity (*M*)	*M*			*M*		

To calculate the molarity of acid in each solution:

$$[\text{acid}] = \frac{(\text{precise molarity of NaOH}) \times (\text{mass of NaOH solution in grams})}{(\text{mass of acid solution in grams}) \times n}$$

The quantity n in the denominator of the above fraction is the number of moles of *titratable* H^+ in each mole of the acid. For HCl, HNO_3 and $HC_2H_3O_2$, n = 1. For H_2SO_4, n = 2, and for $H_3C_6H_5O_7$, n = 3.

Cleaning Up

Clean the plastic cup used for the titration by emptying it into the liquid waste. Rinse with deionized water, and dry if necessary with a paper towel before handing it in. Paper towels go into the solid waste. Wash your hands with soap and water.

Questions for Analysis

1. Show that for a dilute aqueous solution, the number of milligrams (mg) is equal to the number of microliters (μL). What assumption about the solution must you make for this to be true?

2. Why did you *not* have to weigh the empty NaOH pipets before (or after) the titration?

3. Why did you not need to expel air bubbles from the NaOH pipet, or hold it at a vertical angle while titrating?

4. Calculate the %(w/w) of acetic acid in the vinegar.

5. Calculate the %(w/w) of citric acid in the lemon juice.

6. Assume that the true %(w/w) of acetic acid in vinegar is 5.00%. Calculate the percent error in the average value you obtained.