8—Titration of Acids and bases

Name: _____

Date: _____

Section: _____

Objectives

- Reinforce acid-base chemistry principles from chapter 4 in Silberberg
- Standardize a sodium hydroxide solution
- Determine the molarity of an unknown hydrochloric acid solution
- Understand the use of indicators in titrations
- Learn proper pipetting technique
- Learn to titrate a strong acid with a strong base

Pre-Laboratory Requirements

- Read chapter 4.4 in Silberberg
- Watch the instructional videos on pipetting and titrating
- Pre-lab questions (if required by your instructor)
- Laboratory notebook—prepared before lab (if required by your instructor)

Safety Notes

- Eye protection must be worn at all times.
- Hydrochloric acid and sodium hydroxide are caustic and should not come in contact with your skin or clothing. Wear gloves when handling these chemicals. A lab coat or lab apron is recommended.

Discussion

Chemical analysis addresses interesting questions in the world around you and the nature of these questions is boundless. Is on our water safe to drink? Is the correct amount of medication present in the prescription I received from the pharmacist? These are just two examples of the type of questions that are answered every day by analytical chemists.

Titration describes a process where the concentration of an unknown substance is determined by comparing it with a solution of known concentration. The concept that makes titrations possible is finding the equivalence point, i.e., identifying when the quantity of the unknown substance is equal to the quantity of the known substance.

The equivalence point is found in a titration by adding trace amounts of a substance, called an indicator, which turns color when the equivalence point is reached. When a strong acid is titrated with a strong base, or vice versa,

the pH of the solution will be about 7.0 at the equivalence point. Phenolphthalein is the indicator used in this experiment. Phenolphthalein is colorless in acidic solutions and turns pink in alkaline solutions.

This experiment will be done in two parts: (1) preparation and standardization of a 0.1 M sodium hydroxide solution and (2) determination of the acid concentration in an unknown sample provided by the instructor. Standardization is the process of determining concentration in an unknown solution by titrating it with a solution of known concentration.

You should understand the material in chapter 4 of your textbook before beginning this experiment. Pay special attention to the section on "Quantifying Acid-Base Reactions by Titration" and sample problem 4.7, "Finding the Concentration of an Acid from a Titration." (Silberberg, 7th ed., pp 166-168).

Ionic substances dissociate completely when placed in water. Strong acids are ionic substances that form H^+ (or more correctly, H_3O^+) when placed in water and strong bases are ionic substances that form OH^- when dissolved in water. (Note: This is the Arrhenius acid-base definition. You will also learn about the Brønsted-Lowry and Lewis acid-base definitions in CHEM 132).

Hydrochloric acid, HCl, is the strong acid used in this experiment and sodium hydroxide, NaOH, is the strong base. The approximate concentrations of our solutions will be 0.1 M, and the goal of the experiment is to determine the exact concentration of the unknown acid solution provided by the instructor.

Acids contain hydrogen and dissociate in water to give H^+:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Bases contain the hydroxyl group and dissociate in water to give OH^-:

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

An acid (such as HCl) and a base (such as NaOH) react to form water and a salt. In this reaction, we say the acid and the base are **neutralized**:

$$HCl(aq) + NaOH(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l) \text{ (balanced equation)}$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \text{ (net ionic equation)}$$

Throughout this experiment the NaOH solution will be placed in the buret, and the hydrochloric acid solution will always be in the beaker. The phenolphthalein indicator is colorless in acidic solution, and it will turn pink at the equivalence point. A small excess of OH^- is what causes the indicator to change from colorless to pink, which is the titration end point. (*Note: You should stop adding NaOH when the solution in the beaker holds a light pink color for about 10 seconds!).*

Procedure

Part I-A: Preparation of 0.1 M Sodium Hydroxide Solution

Obtain 50.0 mL of 1 M NaOH in a graduated cylinder and pour it into a clean 500 mL polyethylene bottle. Rinse the graduated cylinder with DI water, pouring the rinse water into the polyethylene bottle. Fill the polyethylene bottle to the shoulder with DI water, leaving a small air space in the bottle to allow thorough mixing when the bottle is shaken. Make certain the cap is tightly secured and shake it to make certain the solution is completely mixed. The bottle now contains a sodium hydroxide solution that is approximately 0.1 M NaOH and must be standardized.

Part I-B: Standardization of 0.1 M Sodium Hydroxide Solution

Standardization of the sodium hydroxide solution is accomplished by titrating 25.00 mL of a HCl solution of known concentration. Your instructor will demonstrate the technique for pipetting exactly 25.00 mL HCl into the Erlenmeyer flask.

1. Rinse your buret twice with about 5 mL of the NaOH solution.
2. Fill the buret with the NaOH solution and position it above an Erlenmeyer flask (*Note: be sure there are no air bubbles in the tip of the buret*).
3. Obtain a 25.00 mL pipet and rinse it twice with about 5 mL of the standard HCl solution.
4. Transfer 25.00 mL of the standardized HCl solution into the Erlenmeyer flask.
5. **Add 1-2 drops of phenolphthalein indicator solution into the flask.**
6. Record the **initial** volume of NaOH in the buret before you begin the titration.
7. Begin the titration, swirling the solution in the Erlenmeyer flask as you add NaOH in a drop-wise approach (*Note: Your solution will initially turn pink, and then fade back to colorless when swirled. The pink color will remain longer as you approach the end point of the titration*).
8. Record the **final** volume of NaOH in the buret when the solution in the Erlenmeyer flask remains light pink for about 10 seconds.
9. Repeat this procedure two more times, refilling the NaOH in the buret as needed.

Molarity of standardized HCl: _____

	Trial 1	Trial 2	Trial 3
Volume of HCl delivered from pipet (mL)	_____	_____	_____
Initial buret reading (mL)	_____	_____	_____
Final buret reading (mL)	_____	_____	_____
Volume of NaOH used in titration (mL)	_____	_____	_____
Molarity of NaOH (M)	_____	_____	_____
Average molarity of NaOH (M)		_____	
Average deviation of NaOH (M)		_____	
Relative average deviation of NaOH		_____	

Reported valued for NaOH solution: _____ ± _____ or _____ ± _____ %

Calculations (Note: Use the following relationship to calculate the sodium hydroxide concentration for each trial: (NaOH volume, mL)(**M NaOH**) = (HCl volume, mL)(M HCl).

Part II-A: Titration of an Unknown Acid

The standardized NaOH solution will now be used to determine the concentration of an unknown hydrochloric acid solution. (*Note: You are using the same NaOH solution, but a different acidic solution for this part of the experiment!*).

1. Since your buret is already clean, there is no need to rinse your buret again with the NaOH solution.
2. Fill the buret with the NaOH solution and position it above an Erlenmeyer flask (*Note: be sure there are no air bubbles in the tip of the buret*).
3. Obtain a 25.00 mL pipet and rinse it twice with about 5 mL of the unknown acid solution.
4. Transfer 25.00 mL of the unknown acid solution into the Erlenmeyer flask.
5. Add 1-2 drops of phenolphthalein indicator solution into the flask.
6. Before you begin the titration, be sure to record the **initial** volume of NaOH in the buret.
7. Begin the titration, swirling the solution in the Erlenmeyer flask as you add NaOH in a drop-wise approach approach (*Note: Your solution will initially turn pink, and then fade back to colorless when swirled. The pink color will remain longer as you approach the end point of the titration*).
8. Record the **final** volume of NaOH in the buret when the solution in the Erlenmeyer flask remains light pink color for about 10 seconds.
9. Repeat this procedure two more times, refilling the NaOH in the buret as needed.

Molarity of NaOH (from part I): _____ Unknown number: _____

	Trial 1	Trial 2	Trial 3
Volume of unknown acid delivered from pipet (mL)	_____	_____	_____
Initial buret reading (mL)	_____	_____	_____
Final buret reading (mL)	_____	_____	_____
Volume of NaOH used in titration (mL)	_____	_____	_____
Molarity of unknown acid (M)	_____	_____	_____
Average molarity of unknown acid (M)		_____	
Average deviation of unknown acid (M)		_____	
Relative average deviation of unknown acid		_____	

Reported valued for unknown acid solution: _____ ± _____ or _____ ± _____ %

Calculations (Note: Use the following relationship to calculate the unknown acid concentration for each trial: (NaOH volume, mL)(M NaOH) = (unknown volume, mL)(**M unknown**).

Final Report

Your final report should contain the average calculated value for your unknown acid and the unknown number and the average concentration of your sodium hydroxide solution. Be sure to report the correct number of significant figures and the uncertainty of the unknown acid concentration. Attach a sheet to your report containing your calculations. The calculations do not need to be typed, i.e., they can be written by hand if they are legible.

References

Skoog, D. A.; West, D. M. *Fundamentals of Analytical Chemistry*; Holt, Rinehart and Winston: New York, 1963; pp 341-351.

Sweeder, R. D.; Jeffery, K. A.; A comprehensive general chemistry demonstration. *J. Chem. Ed.,* **2013**, *90*, 96-98. doi:10.1021/ed300367y