Ascorbic Acid Titration of Vitamin C Tablets

This lab will be completed **individually**! Make sure you come prepared!

Introduction

Vitamin C (also known as ascorbic acid, $HC_6H_7O_6$) is a necessary ingredient in the human diet. A deficiency of Vitamin C leads to the disease scurvy, at one time commonly occurring during long sea voyages. British sailors combatted scurvy by carrying limes, rich in Vitamin C, on their voyages, thus engendering the name "limey." Although the Food and Drug Administration recommends a daily intake of 60 mg of Vitamin C, Linus Pauling suggested that amounts of 1-2 grams daily are instrumental in fighting the common cold.

This experiment illustrates how titration can be used to determine the ascorbic acid content of a Vitamin C tablet containing about 500 mg of Vitamin C. First, you will determine the concentration of a sodium hydroxide solution using a standardized solution of sulfuric acid. The mass percentage of ascorbic acid in Vitamin C will then be determined by titrating the Vitamin C samples with the standardized sodium hydroxide solution.

In an acid-base neutralization reaction, an acid reacts with a base to produce water and a salt:

$$\underset{\text{Acid}}{HX\ (aq)} + \underset{\text{base}}{MOH\ (aq)} \rightarrow H_2O\ (l) + \underset{\text{salt}}{MX\ (aq)}$$

The ***protons*** ***(H^+)*** from the acid react with the ***hydroxide ions*** ***(OH^-)*** from the base to form the water. The salt forms by combining the cation from the base and the anion from the acid. Because water is always formed, acids will always react with bases; the solubility of the salt does not determine whether the reaction occurs.

In this experiment, you will determine the molarity (M) of NaOH (aq) using a standardized H_2SO_4 (aq) solution. (A ***standard solution*** has been analyzed, so its concentration is known to a certain degree of accuracy. The H_2SO_4 (aq) solution used in this laboratory was standardized in our stockroom to four significant digits. This concentration will be labeled on the bottle). Write the balanced equation for the reaction that occurs between sodium hydroxide and sulfuric acid on your report sheet.

You will measure out a small volume of sulfuric acid and use a ***buret*** to determine the volume of sodium hydroxide required to completely neutralize the acid. The process of slowly adding one solution to another until the reaction between the two is complete is called a ***titration***.

When carrying out an acid-base neutralization reaction in the laboratory, you observe that most acid solutions and base solutions are colorless, and the resulting water and soluble salt solutions are also colorless. Thus, it is impossible to determine when a reaction has occurred, let alone when it is complete.

To monitor the progress of a neutralization reaction, you will use an ***acid-base indicator***, a solution that changes color depending on the pH (or acid-content) of the solution. One commonly used indicator is ***phenolphthalein***, which is colorless in acidic and neutral solutions and pink in basic (or alkaline) solution. During a titration, the indicator is added to the sample being analyzed. The ***titrant*** is slowly added to the sample until the ***endpoint*** (when the indicator changes color) is reached, signaling that the reaction between the two is complete. Note that phenolphthalein turns pink only when **excess** sodium hydroxide has been added.

If the appropriate indicator has been chosen, the endpoint of the titration (i.e., the color change) will occur when the reaction is complete, or when the acid and base are stoichiometrically equivalent:

moles of acid = moles of base

A Vitamin C tablet contains ascorbic acid, $HC_6H_7O_6$ (aq), as well as binder material that holds the tablet together. The balanced equation for the reaction between ascorbic acid and sodium hydroxide is shown below:

$$HC_6H_7O_6\ (aq) + NaOH\ (aq) \rightarrow H_2O\ (l) + NaC_6H_7O_6\ (aq)$$

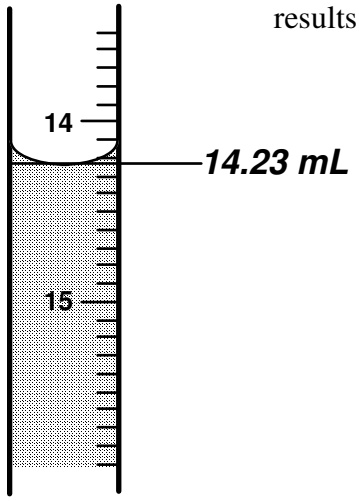
You will titrate each Vitamin C sample with the standardized NaOH solution to determine the mg of ascorbic acid present in each sample.

Techniques:

1) Burets

Burets are used when it is necessary to deliver a liquid to another container and record the exact amount delivered. A buret is marked in milliliters much like a graduated cylinder, except buret markings indicate the number of milliliters **delivered**. This means that **0** (none delivered) is at the top, and the numbers get larger as you go down the buret. The stopcock controls the liquid flow. It is open when parallel to the length of the buret and closed when parallel to the floor.

- **Rinsing and conditioning the buret:** Obtain some deionized water in a small beaker. With the buret over the sink and the stopcock open, pour the water through the buret, letting it drain out the tip into the sink. After the buret is well-drained, close the stopcock and use a small beaker to pour 5-10 mL of the solution to be used (NaOH for this lab) into the buret. Tip the buret sideways and rotate to completely rinse the inside of the buret. Run this solution through the buret tip into your 400 mL waste beaker. This will prevent dilution or contamination and give more accurate results.



- **Filling the buret:** Close the stopcock. Use the beaker of NaOH and a funnel to fill the buret 1 mL above the "0" mark. Place a container under the buret tip and open the stopcock slowly. The buret tip should fill with solution, leaving no air bubbles. If the tip does not fill with solution, ask the instructor for help. Continue to let out solution until the liquid level is at "0" or below.
- **Reading the buret:** Record the volume by noting the bottom of the meniscus. (Be sure that the meniscus is at eye level). If this reading is exactly "0," record 0.00 mL. Otherwise, count the number of markings between each number, and estimate to the nearest 0.01 mL.

Thus, in the example on the previous page, the meniscus is about one-third of the way between 14.2 and 14.3, so the volume is recorded as ***14.23 mL***.

Buret readings are always recorded in mL to <u>2 decimal places</u>.

- Deliver the required volume (usually when you get a color change). To calculate the volume of solution delivered, subtract the initial volume from the final volume.
- ***Always refill your buret before each trial***, so you do not need to refill the buret during a trial. Refilling the buret in the middle of a trial reduces accuracy. When you are finished, empty the buret, and rinse it with DI water, allowing some water to run through the tip.

2) Pipets

You will use a pipet to deliver 10.00 mL of H_2SO_4 to an Erlenmeyer flask. For best results, make sure that the top of the pipet and the bottom of the pipet bulb are dry before use (Note: While a pipet bulb is mentioned in the following text, there are other devices that may be used to draw in solutions to the pipet). You will first condition the pipet (similar to conditioning a buret for a titration). Place the pipet bulb loosely on the top of the pipet, squeeze the bulb, position the tip of the bulb below the liquid level (in your beaker), and slowly release the bulb to draw up a **small amount** of liquid into the pipet, making sure that the tip of the pipet stays below the liquid level. Remove the bulb and quickly slide your index finger or thumb over the top of the pipet. Holding the pipet almost horizontally over a waste beaker in the sink, rotate the pipet in order to coat the inside of the pipet with the solution it will contain. Allow the solution to drain out into your 400 mL waste beaker.

- Replace the bulb, squeeze it, and position the pipet as before. This time, fill the pipet well above the calibration line (etched or marked above the wide center section of the pipet), **taking care not to get liquid into the pipet bulb**. (If this happens, let your instructor know.)
- Slide the bulb off the pipet while quickly sliding your index finger or thumb over the top of the pipet. Move your finger slightly and rotate the pipet to allow the liquid level to drop to the calibration line on the pipet. Then press down harder with your finger and transfer the tip of the pipet into a position over the Erlenmeyer flask.
- Remove your finger and allow most of the liquid to drain out. Then hold the tip of the pipet against the inside of the flask for about 10 seconds to allow more liquid to drain.
- **Do not** try to remove the small amount of liquid remaining in the tip. Pipets are calibrated to retain this amount.

Example Calculations

Titrating an Acid:

Consider the following reaction between hydrochloric acid, HCl(aq), and sodium hydroxide, NaOH (aq),

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

In a titration scenario, the concentration of HCl(aq) is unknown. Using a standardized sodium hydroxide solution with a concentration of 1.020 M, a student titrated 25.00 mL of hydrochloric acid. If 27.14 mL of sodium hydroxide was required to completely neutralize the hydrochloric acid to a faint pink phenolphthalein endpoint, the molarity of the hydrochloric acid is calculated as follows.

The first step in this calculation is recognizing that you are solving for the molarity of hydrochloric acid, which has units of moles per liter and which we can represent as [HCl] (Note: chemists sometimes denote concentration by using brackets as short-hand. For example, "[HCl]" would mean "concentration of HCl"):

$$\text{molarity of HCl} = [\text{HCl}] = \frac{\text{mol HCl}}{\text{L HCl}}$$

Since 25.00 mL of hydrochloric acid was used, convert that to liters (by dividing by 1000, which moves the decimal point to the left three places), and put it in the denominator:

$$\text{molarity of HCl} = [\text{HCl}] = \frac{\text{mol HCl}}{0.02500\ \text{L HCl}}$$

To determine the number of moles of hydrochloric acid, the number of moles of NaOH must first be calculated. ❶ Convert the volume of sodium hydroxide from milliliters to liters then ❷ multiply that by the molarity of sodium hydroxide (given as 1.020 M and shown below as a conversion factor).

By showing the molarity as a fraction (M = mol/L), you can see that the volume units (liters of NaOH) cancel. Now the moles of hydrochloric acid can be calculated in the next step ❸ by using the mole-to-mole ratio between sodium hydroxide and hydrochloric acid found in the balanced chemical equation. The complete calculation to get moles of hydrochloric acid is shown below:

$$\overset{\text{❶}}{\left(\frac{1\ \text{L}}{1000\ \text{mL}}\right)}\overset{\text{❷}}{\left(\frac{1.020\ \text{mol NaOH}}{\text{L of NaOH}}\right)}\overset{\text{❸}}{\left(\frac{1\ \text{mol HCl}}{1\ \text{mol NaOH}}\right)} = \mathbf{0.02768\ mol\ HCl}$$

$$27.14\ \text{mL NaOH}$$

Notice that in the equation above, we know where to place each number (numerator or denominator) by checking if the units cancel.

Finally, we put the number of moles of HCl in the numerator or concentration expression, and the molarity for hydrochloric acid is calculated as follows:

$$\text{molarity of HCl} = [\text{HCl}] = \frac{0.02768\ \text{mol HCl}}{0.02500\ \text{L HCl}} = \mathbf{1.107M\ HCl}$$

Note that there are ***4 significant figures*** in all calculations for this experiment.