EXPERIMENT 15

Determining the Effectiveness of an Antacid

Objective

In this lab you will gain further experience with titrations by observing reactions with hydrochloric acid and commercial antacids, calculate the amount of acid neutralized by commercial antacids, and tabulate and compare results to determine most cost-effective antacid.

Introduction

Double displacement reactions that involve acids and bases are **neutralization reactions**, combining to form salts and water.

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water} \qquad \textbf{(EQ 15.1)}$$

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} \qquad \textbf{(EQ 15.2)}$$

The determination of unknown concentrations of acids or bases can be done using the method known as **titration**. Titrations involve adding one solution whose concentration (of acid or base) is known, the **titrant**, usually from a buret to a specific volume of an another solution of unknown concentration (acid or base), the **analyte**. One mole of H^+ ion neutralizes one mole of OH^-. As the reaction proceeds, the pH of the solution changes. When all H^+ ions have been neutralized, this is the titration's **equivalence point.** At this point the addition of the basic solution will have a dramatic effect on pH. This point may be detected visually when a color change occurs by using an appropriate indicator, or by measuring the change in pH with a pH meter. A chemical **indicator** which changes color at various pHs, can be used to find the **end point**, which should be near the equivalence point. For example, the indicator phenolphthalein is a pinkish-red in basic solutions between pH 8.2-10.0 and colorless in acidic solutions changes colors. There is often a slight difference between the change in the indicator color and the actual equivalence point of the titration. This is an indeterminate error (i.e. systematic error). To minimize this error an indicator with an end point very close to the equivalence point of the reaction should be used.

Back titration is a method of indirect titration, where the concentration of the analyte is determined by reacting it with a known number of moles of excess reagent. The excess reagent is then neutralized by titrating it against a second reagent of known concentration. The concentration of the analyte in the original solution can be found based on the amount of reagent consumed.

EXAMPLE 15.1

A student places 100.00 mL of 0.1000 *M* HCl solution in a flask with a crushed antacid table and phenolphthalein indicator. After the reaction has gone to completion, he titrates the excess acid with 0.1000 *M* NaOH solution. The initial reading of the buret is 0.40 mL. At the end point, the final reading of the buret is 27.15 mL. The student wants to determine the number of H^+ moles that were neutralized by the antacid.

First, calculate the volume of base required to neutralize the acid.

$$\text{final volume} - \text{initial volume} = \text{volume of base required to neutralize acid}$$
$$27.15\ \text{mL} - 0.40\ \text{mL} = 26.75\ \text{mL required to neutralize acid} \quad \textbf{(EQ 15.3)}$$

Second, calculate how many moles of base were used to neutralize the acid.

$$\text{molarity of base} \times \text{L of base used to neutralize acid} = \text{moles of base used}$$
$$0.02675\ \text{L} \times \frac{0.1000\ \text{mol NaOH}}{1\ \text{L NaOH soln}} = 0.002675\ \text{mol NaOH} \quad \textbf{(EQ 15.4)}$$

Third, calculate the number of moles of acid neutralized by the added NaOH. Since this is a monoprotic acid it is a one-to-one reaction. Therefore, one mole of base neutralizes one mole of acid.

$$0.002675\ \text{mol NaOH} \times \frac{1\ \text{mol HCl}}{1\ \text{mol NaOH}} = 0.002675\ \text{mol HCl} \quad \textbf{(EQ 15.5)}$$

Fourth, determine how many moles of acid were in the flask prior to titration.

$$0.10000\ \text{L HCl} \times \frac{0.1000\ \text{mol HCl}}{1\ \text{L HCl soln}} = 0.01000\ \text{mol HCl} \quad \textbf{(EQ 15.6)}$$

Fifth, calculate number of moles of acid that the antacid tablet neutralized. The number neutralized is the difference between the number that was added initially and that left over.

$$\text{mol HCl before neutralization} - \text{mol HCl neutralized by titration} = \text{mol HCl neutralized by antacid} \quad \textbf{(EQ 15.7)}$$

$$0.01000\ \text{mol HCl} - 0.002675\ \text{mol HCl} = 0.007325\ \text{mol HCl} = 0.00733\ \text{mol HCl neutralized} \quad \textbf{(EQ 15.8)}$$

A back titration is useful if the end point of the revise titration is easier to identify than the end point of the normal titration. They are also useful if the reaction between the analyte and the titrate is very slow.

Hydrochloric acid is the acid present in our stomachs. Many of us are plagued with a condition called heartburn. Heartburn is a result of acid reflux into the esophagus. Commercial **antacids** are effective in neutralizing stomach acid. How effective an antacid may be is determined by the amount of acid neutralized. In this experiment you will make this determination. The mixing of hydrochloric acid and a commercial antacid in an Erlenmeyer flask mimics what goes on in the stomach. The antacid is made up of active and inactive ingredients. The base component of the antacid is the active ingredient, calcium carbonate, sodium bicarbonate, and magnesium hydroxide are common examples:

$$2\ HCl_{(aq)} + CaCO_{3\,(s)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)} \quad \textbf{(EQ 15.9)}$$

$$HCl_{(aq)} + NaHCO_{3\,(s)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2\,(g)} \quad \textbf{(EQ 15.10)}$$

$$2\,HCl_{(aq)} + Mg(OH)_{2\,(s)} \rightarrow MgCl_{2\,(aq)} + 2\,H_2O_{(l)} \quad \textbf{(EQ 15.11)}$$

Once the antacid has exhausted its neutralizing capabilities there will be acid left in the flask. You will determine and compare the neutralizing capabilities of several commercial antacids. The more effective the antacid, the more moles of acid it can neutralize. You will titrate the solution with a known concentration of base to figure out how many moles of acid are left in the flask after the antacid has been used up. This is an example of a back titration.

After gathering and analyzing your data, you will be able to do determine mass effectiveness:

$$\text{mass effectivenes} = \frac{\text{mol HCl neutralized by antacid}}{\text{g antacid tablet}} \quad \textbf{(EQ 15.12)}$$

and cost effectiveness of each antacid:

$$\text{cost effectivenes} = \frac{\text{mol HCl neutralized by antacid}}{\text{cost of tablet}} \quad \textbf{(EQ 15.13)}$$

You will also determine the mass of base per tablet and compare it to the manufacturer's values. Using this information you will determine the "best" antacid from the ones supplied.

Procedure

Part A: Sodium Hydroxide Standardization

1. In a beaker prepare about 400 mL of a 0.1 M sodium hydroxide solution from the pellets provided. Record the mass of the sodium hydroxide to as many decimal places as you can. Be sure to stir well.

NOTE: Notice how the numbers on the balance are increasing! Sodium hydroxide is hygroscopic.

2. Obtain a volumetric pipet and buret from the stockroom. Be sure to throughly clean and condition them.
3. Carefully pipet 10.00 mL of the 1.00 *M* HCl into the Erlenmeyer flask.
4. Add an appropriate indicator, such as bromothymol blue, to your solution.
5. Fill the buret with the NaOH solution. Record the initial volume to 2 decimal places.
6. Titrate the HCl solution by adding the NaOH solution drop-wise to the flask that contains your acid and the indicator. Gently swirl the flask while titrating. As you approach the end point, be sure to rinse off the tip of the buret and the sides of the flask with deionized water. When you see the indicator permanently change color that lasts at least 60 seconds, you have reached the end point.
7. Record the final volume of the base solution.
8. Repeat the procedure four more times. Your volume of base added should be within 0.05 mL of each other.

Part B: Antacid Titration

1. Obtain one tablet of antacid. Weigh the tablet and record the weight to the nearest 0.001 g.
2. Crush the tablet with the mortar and pestle.
3. Add the crushed tablet quantitatively to an Erlenmeyer flask. Use a powder funnel or weigh paper to transfer the powdered tablet to the flask and rinse the mortar, pestle, and funnel with several aliquots of DI water to be sure that all of the antacid has been transferred to the Erlenmeyer flask.
4. Obtain a volumetric pipet and buret from the stockroom. Be sure to throughly clean and condition them. If you have not already done so.
5. Carefully pipet 10.00 mL of the 1.00 *M* HCl into the Erlenmeyer flask with the crushed tablet and swirl. Gently heat the mixture to help dissolve the tablet and any inert ingredients, which may be in the antacid.
6. If the tablet does not dissolve add an additional 10.00 mL of acid.
7. Test with litmus paper. If still basic continue to add aliquots of HCl until the solution is acidic to litmus.
8. Add an appropriate indicator, such as phenolphthalein, to your now acidic solution.
9. Obtain about 150 mL of NaOH in a beaker. Fill the buret with the NaOH solution. Record the initial volume to 2 decimal places.
10. Titrate the antacid solution by adding the NaOH solution drop-wise to the flask that contains your antacid and the indicator. Gently swirl the flask while titrating. As you approach the end point, be sure to rinse off the tip of the buret and the sides of the flask with deionized water.

When you see the indicator permanently change color that lasts at least 60 seconds, you have reached the end point.

11. Record the final volume of the base solution.
12. Repeat the procedure four more times, remembering to use the same brand of antacid.
13. Titrate with at least one other antacid or the number specified by your instructor.

Partner Variation

1. Work with a partner and as a group titrate four different antacids and compare the antacids in terms of effectiveness.

Results and Calculations

1. Complete Table 15.2 to determined the concentration of the NaOH solution and the standard deviation.
2. Complete Table 15.4 and Table 15.5 to analyze the antacid tablets. Be sure to show one calculation of each type.
3. What does your standard deviation tell you about your results?
4. What does the percent deviation tell you about your results?
5. Identify the best antacid based on your analysis of your data.

Report

Copy Table 15.2 , Table 15.3 , and Table 15.4 into your lab notebook to use during the experiment. Complete Table 15.2 , Table 15.4 , and use Table 15.5 to analyze your data.

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TABLE 15.2 Standardization of NaOH Solution

	Trial #1	Trial #2	Trial #3	Trial #4
Molarity of HCl solution				
Volume of analyte				
Initial buret reading				
Final buret reading				
Volume of titrant				
Molarity of NaOH solution				
Average Molarity of NaOH solution				
Standard Deviation				

TABLE 15.3 Reagent Data

Brand of antacid	
Cost of antacid	per tablets
Amount of base per tablet	
Concentration of HCl solution	
Concentration of NaOH solution	

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TABLE 15.4 Antacid Back-Titration Data using Standardized NaOH Solution

	Trial #1	Trial #2	Trial #3	Trial #4
Mass of flask				
Mass of flask and tablet				
Mass of tablet				
Volume HCl added				
Initial buret reading				
Final buret reading				
Volume of titrant				

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TABLE 15.5 Antacid Analysis

	Trial #1	Trial #2	Trial #3	Trial #4
Number of moles of HCl initially added				
Number of mol NaOH from buret				
Number of moles of base from tablet (i.e. moles of HCl neutralized by antacid tablet)				
Number of moles HCl neutralized per gram tablet				
Average number moles of acid neutralized per tablet				
Standard deviation				
Average number of moles neutralized per gram tablet				
Standard deviation				
Cost of antacid per tablet				
$\frac{\text{average number of moles HCl neutralized}}{\text{cost of tablet}}$				
average moles of base/tablet				
average mass of base/tablet				
% deviation				

Post Lab Questions

1. Which of the antacid tablets analyzed by your lab section is the most cost effective?

2. Calcium dietary supplements are composed primarily of calcium carbonate which is also the main component of marble.
   a. Briefly explain why calcium carbonate would be a good calcium ion source for those suffering from chronic heartburn.

   b. Marble statues erode when exposed to acid rain. Give a brief explanation for this erosion; include a chemical equation.

3. How does the dye affect your titration? Propose some modifications that might be made to the procedure if you are trying to analyze a pink colored tablet.

4. Carbonate functions as a buffer. **Buffers** are weak acids or bases, which are capable of minimizing pH changes. How might this affect the results of your titrations?

5. A student analyzed an antacid tablet from a bottle of 50 tablets purchased at a discount store for $1.00. The mass of the tablet was 1.470 g. After adding 25.00 mL of 0.7989 *M* HCl solution to the tablet, the student back-titrated the excess HCl with 3.55 mL of 1.017 *M* NaOH solution.

   a. Calculate the number of moles of stomach acid neutralized by one tablet.

   b. Calculate the mass effectiveness of the tablet.

   c. Calculate the cost effectiveness of one tablet.