EXPERIMENT 10
Titration of a Commercial Antacid

Objective:

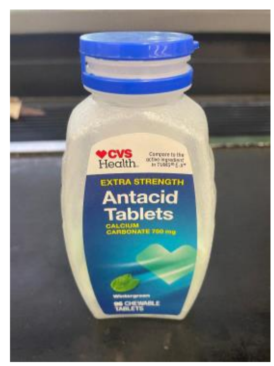
To quantify the neutralizing capacity of a single TUMS tablet against stomach acid.

Equipment and Materials:

Buret, Support stand, Double Buret Clamp, Hotplate with magnetic stirrer, 25 mL volumetric pipets with pipet pumps, 250mL Erlenmeyer flask, 100-mL beaker, Mortar and pestle with splint, Wintergreen TUMS (Calcium Carbonate) 750mg, ~0.2M Sodium Hydroxide, ~0.4M Hydrochloric Acid, and Phenolphthalein

**** * * GOGGLES ARE REQUIRED FOR THIS EXPERIMENT. * * ****

Introduction:



The parietal cells in the stomach secrete hydrochloric acid (HCl) at a concentration of roughly 0.16 M. The flow of HCl increases when food enters the stomach. If you eat or drink too much, you may develop heartburn or indigestion. Antacids, such as TUMS are used to neutralize this excess acid. The active ingredient in TUMS is calcium carbonate, $CaCO_3$, a base. There are also other ingredients, such as binders present in each tablet. On average, about 40% of a tablet's mass is from calcium carbonate.

(1) Hydrochloric acid is a strong acid and forms the hydrogen ion in water:

$HCl\ (aq) \rightarrow \underbrace{H^+\ (aq)}_{\text{Hydrogen ion}} + Cl^-\ (aq)$

(2) The hydrogen ion is neutralized by calcium carbonate, forming the bicarbonate ion:

$CaCO_3\ (s) + H^+\ (aq) \rightarrow Ca^{2+}\ (aq) + \underbrace{HCO_3^-\ (aq)}_{\text{Bicarbonate}}$

(3) Bicarbonate then reacts with another hydrogen ion to form carbonic acid:

$HCO_3^-\ (aq) + H^+\ (aq) \rightarrow \underbrace{H_2CO_3\ (aq)}_{\text{Carbonic acid}}$

(4) Carbonic acid then decomposes into carbon dioxide and water, producing bubbles of gas:

$H_2CO_3\ (aq) \rightarrow CO_2\ (g) + H_2O\ (l)$

The net reaction is:
$2\ HCl\ (aq) + CaCO_3\ (s) \rightarrow Ca^{2+}\ (aq) + CO_2\ (g) + H_2O\ (l)$

This means that each calcium carbonate formula unit can neutralize two hydrochloric acid molecules in water. Calcium ions are generally safe as is water. The carbon dioxide gas will escape the stomach through the mouth, resulting in a reduction of discomfort.

Titration is a common laboratory technique in chemistry used to determine the concentration of an unknown solution by gradually adding a solution of known concentration (called the **titrant**) until a chemical reaction is complete. This point is usually indicated by a color change, thanks to a substance called an **indicator**.

To determine the effectiveness of an antacid in neutralizing acid, we will begin by dissolving the tablet in an excess amount of hydrochloric acid (HCl) of known concentration. A portion of the HCl will be neutralized by the antacid, while the remainder will stay unreacted. To determine the amount of unreacted acid, we will perform a titration using a sodium hydroxide (NaOH) solution. This allows us to calculate the amount of acid that reacted with the antacid by difference. This type of analysis is known as **back-titration**. Since the reactions involved are reversible, carbon dioxide (CO_2) produced during the reaction can dissolve in water and form carbonic acid. This additional acid could react with the NaOH during titration, leading to inaccurate results. To prevent this, we will boil the mixture to drive off CO_2 from the solution before titration.

Procedure:

You will work in pairs, but each student should perform one of the titrations. Rinse all the glassware from your drawer that you plan to use.

1. Hydrochloric Acid

In the fume hood, obtain approximately 150 mL of hydrochloric acid (HCl) solution into a clean, dry 250-mL beaker and bring it to your workstation. Be sure to write the concentration of HCl here so that you can use it in your calculations: _______________ M.

2. Transferring HCl to Erlenmeyer Flask

Using a 25-mL volumetric pipet, transfer two full portions (totaling 50.00 mL) of the HCl solution into a clean 125-mL clean, dry Erlenmeyer flask. Your instructor will demonstrate how to use the volumetric pipet and pipet pump.

3. Preparing the Antacid

Take one antacid tablet and crush it thoroughly with a mortar and pestle.

4. Reaction with Antacid

Using a clean splint, carefully transfer as much of the powdered antacid tablet as possible into the HCl solution contained in the Erlenmeyer flask. The acid-antacid mixture should turn green due to dyes in the tablet. Place a Teflon-coated magnetic stir bar in the acid-antacid mixture and set the flask on a magnetic stirrer with heating. Turn on the stirrer to mix the contents thoroughly and apply heat until the mixture reaches a gentle boil. Maintain a gentle boil for 2 minutes. After gently boiling for 2 minutes, turn off the heat. There should be no visible solid residue remaining. If necessary, stir a little longer to assist dissolution.

** While you're waiting for the mixture to boil, you may proceed to step 5,6 and 7. **

5. Sodium Hydroxide Solution

In the fume hood, obtain approximately 60 mL of sodium hydroxide (NaOH) solution into a clean, dry 100-mL beaker and bring it to your workstation. Write the concentration of the solution here so you have the number for your calculations: _____________M.

6. Preparing the Buret

Rinse the 50-ml buret first with tap water, then with distilled water, and finally pour 2–3 mL of NaOH solution through the buret. Your instructor will show you how to prepare the buret.

7. Filling the Buret

Position the buret so when you fill the buret, the NaOH solution remains below eye level. Using a funnel, fill the buret with the remaining NaOH solution, making sure the bottom valve is closed. The liquid level should rise slightly above the zero mark. Open the valve briefly to allow a few milliliters to drain through—this removes air bubbles. Once air is cleared, adjust the liquid level to exactly 0.00 mL, which will be your "initial volume". Record this value on your data sheet (page 4).

8. Adding Indicator

Add two drops of phenolphthalein indicator to the acid-antacid mixture in the Erlenmeyer flask.

9. Titration

Begin the titration by slowly adding NaOH solution from the buret, drop by drop, to the acid-antacid mixture in the Erlenmeyer flask. Allow the magnetic stirrer to mix. Titrate until the solution turns and remains a light violet for 30 seconds. (Normally phenolphthalein turns pink in basic solution). This is the end-point of the titration. Measure the level of NaOH solution in the buret. This will be your "final volume". Write the value in the data sheet (page 4).

10. Repeat the steps from 2-9. You will have data for at least two trials.

11. Dispose of waste solutions in the designated waste containers. Unplug the magnetic stirrer. Clean your bench top. Rinse all glassware and return it to your drawer. Return and place all equipment to its proper place.

Data table:

	Trial 1	Trial 2
Amount of HCl in milliliters	mL	mL
Amount of HCl in liters	L	L
Final volume in buret in milliliters	mL	mL
Initial volume in buret in milliliters	mL	mL
Volume of NaOH dispensed in milliliters	mL	mL
Volume of NaOH dispensed in liters	L	L
Average volume of NaOH in liters	L	

Note: volume in liters = volume in milliliters ÷ 1000 mL/L

Calculation and Results:

For each calculation be sure to show your reasoning in detail. Pay attention to units and significant digits.

1. Using the average volume of the NaOH (in liters), calculate the number of moles of NaOH that were added from the buret. The equation for molarity is Molarity = moles/volume. You wrote the molarity in step 5 of the procedure.

Solving for moles we get: Moles = (Molarity) x (Volume)

Moles NaOH used in titration = ____________________ mol NaOH

***This number is equal to the number of moles of HCl that were neutralized by the NaOH since they react in a 1:1 ratio: *HCl (aq) + NaOH (aq)* $\rightarrow$ *NaCl (aq) +* H_2O *(l)*

2. Calculate the number of moles of HCl that you started with. This is referring to the 50.00 mL of HCl you added with the pipet and the molarity of HCl solution you used. Remember that the volume here needs to be in liters, not milliliters. You wrote the molarity in step 1 of the procedure.

Moles = (molarity) x (volume)

Moles of HCl at start = ____________________________ mol HCl

3. Calculate the number of moles of HCl that were neutralized by the antacid. This can be found by taking the number of moles of acid that you began with (from calculation 2) and subtracting the number of moles HCl that were in excess (from calculation 1).

Moles HCl neutralized by antacid = __________________ mol HCl

4. Use the number of moles of HCl neutralized by the tablet (calculation 3) to determine the mass of the calcium carbonate in the tablet. The ratio of $CaCO_3$ to HCl is 1:2.

$2\ HCl\ (aq) + CaCO_3\ (s) \rightarrow CaCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$

The molar mass of $CaCO_3$ is 100 g/mol.

$mg\ CaCO_3 = (mol\ HCl\ neutralized)\left(\frac{1\ mol\ CaCO3}{2\ mol\ HCl}\right)\left(\frac{100\ g\ CaCO3}{1\ mol\ CaCO3}\right)\left(\frac{1000\ mg}{1\ g}\right)$

mg $CaCO_3$ = ______________________ mg

Data Analysis:

Compare your experimental value to the known value (750 mg) and calculate a percent error.

$$\%\, error = \frac{(experimental - known)}{known} x\ 100\%$$

% error = ________________

Discussion/Conclusions:

1. How can you determine when the titration has reached its endpoint?

2. What could have caused inaccuracies or errors during the experiment?

3. Did the tablets show consistency in their antacid content?