

Experiment #11: Titration of a Commercial Antacid

Objective: To measure the quantity of stomach acid that can be neutralized by a tablet of Tums.

Safety glasses are required for this experiment.

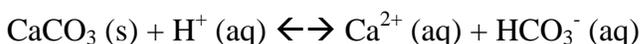
Textbook Reference: pp 455, 465-469

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, a 1.3 gram tablet contains 0.5 g of calcium carbonate.

HCl is neutralized by calcium carbonate as illustrated below:



To determine the ability of Tums to neutralize acid, we are first going to dissolve the tablet in an excess amount of acid of known concentration. Some of the HCl will be neutralized by the carbonate, but there will be some remaining. We will then perform a titration with NaOH to figure out the amount of excess acid. Then, from this, we can calculate how much acid reacted with the antacid. This method of analysis is called a back-titration.

The reactions above are reversible, which means that CO_2 dissolved in water will produce some carbonic acid. This acid will react with the NaOH we are titrating and give us inaccurate results. Therefore it is important to boil the solution when the carbonate reacts with acid, to remove CO_2 as a gas.

Procedure:

You will work in pairs, but each student should perform one of the titrations.

1. In a 125 mL Erlenmeyer flask, pipet in exactly 50.00 mL of _____ M HCl. You will be using a 25.00 mL pipet.
2. Take 1 tablet of Tums. Crush it up into small pieces.
3. Dissolve the Tums in the acid in the Erlenmeyer flask and boil this solution for 2 minutes.
4. Add _____ M NaOH to the buret. Your instructor will show you how to prepare and read the buret.

5. After boiling, add 3 drops of phenolphthalein indicator to the Erlenmeyer flask. An acidic solution would normally be clear at this point, however since we are using wintergreen Tums, it will be green.
6. You are now ready to titrate. Slowly add the NaOH from the buret while constantly swirling the Erlenmeyer flask. Titrate until the solution remains violet for 30 seconds. (Normally phenolphthalein turns pink in basic solution). This is the end-point of the titration. Make sure you record the final volume of the buret.

Data:

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

Calculations:

- 1) Using the average volume of the 2 titrations, calculate the number of moles of NaOH that were added from the buret.
 The equation for molarity is $\text{Molarity} = \text{moles}/\text{volume}$
 Solving for moles we get:
 $\text{Moles} = (\text{Molarity})(\text{Volume})$

*****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.**

- 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)

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 #2) and subtracting the number of moles HCl that were in excess (from #1).

4) Given that the molarity of stomach acid is approximately 0.16 M, calculate the volume of stomach acid that could be neutralized by 1 tablet of Tums.

$$\text{Volume} = \text{moles/molarity}$$

5) To determine a percent error for your titration, use the number of moles of HCl neutralized by the tablet to determine the mass of the calcium carbonate in the tablet. The ratio of CaCO₃ to HCl is 1:2. The molar mass of CaCO₃ is 100 g/mol.

$$\text{mg CaCO}_3 = \text{mols HCl neutralized} \times \frac{1 \text{ mol CaCO}_3}{2 \text{ mol HCl}} \times \frac{100 \text{ g}}{1 \text{ mol}} \times \frac{1000 \text{ mg}}{1 \text{ g}} =$$

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

$$\% \text{ error} = \frac{(\text{measured} - \text{known})}{\text{known}} \times 100\%$$