The Acid Neutralizing Potential of Antacids

Lab #2, Chem 36

Spring 2009

Introduction

One of the causes of indigestion is the secretion of excess gastric juice, 0.25% HCl, and a sensitization of the stomach lining to this acid. The reasons for the secretion of stomach acid are many, and include eating spicy or greasy foods, eating too much too fast, drinking too much, or a number of other poor (but frequently enjoyable) eating habits. In order to alleviate the discomfort of their fellow men and women, and to make a profit, a number of pharmaceutical companies have produced remedies which combat this discomfort. These remedies work because they contain weak bases such as, sodium bicarbonate, magnesium carbonate, aluminum hydroxide, etc. All of the these compounds are able to neutralize the excess stomach acid without causing further harm.

In this experiment the acid neutralizing potential of a commercial antacid will be determined by the back titration method. Weak or unreactive compounds, such as the bases in antacids, are often difficult to analyze by a direct titration because they react slowly and give end points that are difficult to observe. In a back titration an excess of a strong standardized reagent, HCl in this experiment, is added to the substance being analyzed. After waiting a few minutes for the reaction to be completed, the excess strong reagent is titrated. From the titration the mol of strong reagent in excess is determined. Subtracting the mol of reagent in excess from the total mol of strong reagent initially added, the mol of the compound being analyzed is obtained.

After determining the relative strength of your antacid compare your value to those of other students doing different products. Which antacid provides the greatest acid neutralizing potential?

Experimental

1. <u>Preparation of approximately 0.5 M HCl</u>. Add 10 mL of concentrated HCl to 240 mL of distilled water and mix well. Since this solution has to be standardized, there is no point in being overly precise about adding exactly 10 mL of HCl or exactly 240 mL of water. Do not pipette the concentrated HCl.

2. <u>Preparation of a standardized NaOH solution</u>. A previously standardized solution of NaOH will be available in the laboratory. Record its exact concentration and take approximately 10 mL in a small beaker. Carefully pipette 5.00 mL of this solution into a 250 mL volumetric flask. Dilute the flask to the mark with distilled water and mix thoroughly. You now have a solution of NaOH with a precisely known molarity.

3. <u>Standardization of the HCl solution with NaOH</u>. Obtain two burets from the stockroom. Fill one buret with the HCl from step 1 and the other with the NaOH from step 2. Run 5.00 mL of the HCl solution into an Erlenmeyer flask. Add about 30 mL of distilled water, 2-4 drops of phenolphthalein indicator and swirl to mix. Titrate the HCl solution with the standardized NaOH until the first permanent faint pink color persists. **Do three trials**.

Calculate the molarity for each of the HCl trails and check to make sure the trials are within 5% of each other. If the molarities from the three trials are more than 5% apart, do a fourth trial. The percent difference between the trials may be calculated by the method shown below. Average the results from the three good trials.

$$\frac{2\times\left[\left(M\ HCl\ high\right)-\left(M\ HCl\ low\right)\right]}{\left(M\ HCl\ high\right)+\left(M\ HCl\ low\right)}\times 100$$

Waste

The neutralized solution may be poured down the drain.

4. <u>The acid neutralizing capacity of an antacid sample</u>.

a. Take 4 to 5 antacid tablets, weigh them and then grind them up in a mortar. Accurately weigh an antacid sample of between 1.300 to 1.500 grams and transfer it to a clean Erlenmeyer flask.
b. Run 25 mL of the standardized HCl into the flask and allow the mixture to react for 5 minutes. The more finely the tablets are ground-up the faster the active portion of the tablet will dissolve. The mixture will still be cloudy as the binder is insoluble. Wash the sides of the flask with 5-10 mL of distilled water and add 3-5 drops of bromophenol blue indicator. If the solution is blue and not yellow at this time run an additional 5 mL of the HCl solution into the mixture. Keep adding HCl, 5 mL at a time, until the solution turns yellow.
c. After the color of the antacid mixture is yellow, titrate the mixture using the standardized NaOH to its blue end point. **Do three trials**.
d. Before leaving, be sure to obtain the cost of the tablets as it is needed for the calculations.

Waste

The neutralized solution from each titration may be poured down the drain. Mix the left over HCl and NaOH together and pour the neutralized solution down the drain.

Calculations

1. Calculate the molarity of the NaOH solution after dilution.
2. Determine the molarity of the HCl for each of the trails and average the results to obtain the standard value of the HCl solution.
3. Calculate the total moles of HCl added to each antacid (AA) sample by multiplying the volume of HCl times the molarity of the HCl.
4. Determine the moles of HCl in excess by multiplying the volume of NaOH titrated into the AA sample by the molarity of the NaOH and then converting to moles of HCl.
5. By subtracting the moles of HCl in excess from the total moles of HCl added, the moles of HCl that reacted with the AA can be determined.
6. Assuming that one mole of HCl reacts with one mole of AA, the moles of AA per gram of AA sample can be determined.
7. Calculate the mL of stomach acid "consumed" or neutralized per tablet and per gram of antacid. Assume stomach acid is 0.25% HCl by mass and has a density of 1.19 g/mL.
8. Calculate how much a mole of AA costs per penny. A recent catalog showed pure NaOH selling at approximately $4.00 per kg. Is the base in the antacid sample more or less expensive then NaOH?
9. Explain why bromphenol blue (pK_a 3.8) was used for the titration of the antacid, rather than phenolphthalein (pK_a 9.0).