To determine the neutralizing effectiveness per gram of a commercial antacid

Various commercial antacids claim to be the “most effective” for relieving acid indigestion. All antacids, regardless of their claims or effectiveness, have one purpose: to neutralize the excess hydrogen ion in the stomach to relieve acid indigestion.

The acidity (and basicity) of an aqueous solution is often expressed in terms of its pH. At 25°C, acidic solutions have a pH less than 7, with the lower values being the more acidic; basic solutions have a pH greater than 7, with the higher values being more basic.

The pH of the “gastric juice” in the stomach ranges from 1.0 to 2.0. This acid, primarily hydrochloric acid, is necessary for the digestion of foods. Acid is continually secreted while eating; consequently, overeating can lead to an excess of stomach acid, leading to acid indigestion and a pH less than one. An excess of acid can, on occasion, cause an irritation of the stomach lining, particularly the upper intestinal track, causing “heartburn.” An antacid reacts with the hydroxion ion to relieve the symptoms. Excessive use of antacids can cause the stomach to have a pH greater than 2, which stimulates the stomach to excrete additional acid, a potentially dangerous condition.

The most common bases used for over the counter antacids are:

- aluminum hydroxide, Al(OH)₃
- magnesium hydroxide, Mg(OH)₂
- calcium carbonate, CaCO₃
- sodium bicarbonate, NaHCO₃
- potassium bicarbonate, KHCO₃

Milk of magnesia (Figure 26.1), an aqueous suspension of magnesium hydroxide, Mg(OH)₂, and sodium bicarbonate, NaHCO₃, commonly called baking soda, are simple antacids (and thus, bases) that neutralize hydroxion ion, H₂O⁺:

\[
\text{Mg(OH)}_2(s) + 2 \text{H}_2\text{O}^+(aq) \rightarrow \text{Mg}^{2+}(aq) + 4 \text{H}_2\text{O}(l) \quad (26.1)
\]

\[
\text{NaHCO}_3(aq) + \text{H}_2\text{O}^+(aq) \rightarrow \text{Na}^+(aq) + \text{CO}_2(g) + 2 \text{H}_2\text{O}(l) \quad (26.2)
\]

The release of carbon dioxide gas from the action of sodium bicarbonate on hydronium ion (Equation 26.2) causes one to “belch.”

To decrease the possibility of the stomach becoming too basic from the antacid, buffers are added as part of the formulation of some antacids. The more common, “faster relief” commercial antacids that buffer excess acid in the stomach are those

All antacids, as weak bases, reduce the acidity of the stomach.
Buffer: a condition that resists changes in the acidity or basicity of an aqueous system.

containing calcium carbonate, CaCO₃, and/or sodium bicarbonate. A HCO₃⁻/CO₃²⁻ buffer system¹ is established in the stomach with these antacids.

\[
\text{CO}_3^{2-}(aq) + \text{H}_3\text{O}^+(aq) \rightarrow \text{HCO}_3^-(aq) + \text{H}_2\text{O}(l)
\]  
(26.3)

\[
\text{HCO}_3^-(aq) + \text{H}_3\text{O}^+(aq) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(l)
\]  
(26.4)

Rolaids (Figure 26.2), containing sodium aluminum dihydroxycarbonate, but labeled dihydroxyaluminum sodium carbonate, NaAl(OH)₂CO₃, is a combination antacid that neutralizes hydronium ion, releasing bicarbonate ion, HCO₃⁻:

\[
\text{NaAl(OH)}_2\text{CO}_3(aq) + 3 \text{H}_3\text{O}^+(aq) \rightarrow \text{Na}^+(aq) + \text{Al}^{3+}(aq) + \text{HCO}_3^-(aq) + 5 \text{H}_2\text{O}(l)
\]  
(26.5)

<table>
<thead>
<tr>
<th>Table 26.1 Common Antacids</th>
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<tbody>
<tr>
<td>Principal Active Ingredient</td>
</tr>
<tr>
<td>NaHCO₃</td>
</tr>
<tr>
<td>KHCO₃</td>
</tr>
<tr>
<td>Mg(OH)₂</td>
</tr>
<tr>
<td>NaAl(OH)₂CO₃</td>
</tr>
<tr>
<td>CaCO₃</td>
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<tr>
<td>Al(OH)₃/MgCO₃</td>
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<td>Al(OH)₃/Mg(OH)₂</td>
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In this experiment, the “neutralizing powers” of several antacids are determined using a strong acid–strong base titration. To obtain the quantitative data for the analysis, which requires a well-defined endpoint in the titration, the buffer action must be destroyed.

The buffer is destroyed when an excess of hydrochloric acid, HCl, is added to the antacid solution; this addition drives the reaction in Equation 26.4 to the right. The antacid solution is heated to remove carbon dioxide. The excess HCl is then titrated with a standardized sodium hydroxide, NaOH, solution.² This analytical technique is referred to as a “back” titration.

¹A buffer system resists large changes in the acidity of a solution. To analyze for the amount of antacid in this experiment, we want to remove this buffering property to determine the total effectiveness of the antacid. To do this, we “swamp” the system with an excess of strong acid, HCl, and then titrate with a strong base, NaOH, to determine the amount of strong acid that was not neutralized by the antacid (base).

²A standardized NaOH solution is one in which the concentration of NaOH has been very carefully determined.
The number of moles of base in the antacid of the commercial sample plus the number of moles of NaOH used in the titration equals the number of moles of HCl added to the original antacid sample:

\[
\text{moles}_{\text{antacid}} + \text{moles}_{\text{NaOH}} = \text{moles}_{\text{HCl}}
\]

(26.6)

A rearrangement of the equation provides the moles of base in the antacid in the sample:

\[
\text{moles}_{\text{antacid}} = \text{moles}_{\text{HCl}} - \text{moles}_{\text{NaOH}}
\]

(26.7)

The moles of base in the antacid per gram of antacid provides the data required for a comparison of the antacid effectiveness of commercial antacids.

**Procedure Overview:** The amount of base in an antacid sample is determined. The sample is dissolved, the buffer components of the antacid are destroyed with an excess of standardized HCl solution, and the unreacted HCl is titrated with a standard NaOH solution.

At least two analyses should be completed per antacid if two antacids are to be analyzed to compare their neutralizing powers. If only one antacid is to be analyzed, complete three trials.

1. **Determine the Mass of Antacid for Analysis.** If your antacid is a tablet, pulverize and/or grind the antacid tablet with a mortar and pestle. Measure (±0.001 g) no more than 0.2 g of the pulverized commercial antacid tablet (or 0.2 g of a liquid antacid) in a 250-mL Erlenmeyer flask having a known mass.

2. **Prepare the Antacid for Analysis.** Pipet 25.0 mL of standardized 0.1 M HCl (stomach acid equivalent) into the flask and swirl. Record the actual molar concentration of the HCl on the Report Sheet. Heat the solution to a gentle boil and maintain the heat for 1 minute to remove dissolved CO₂. Add 4–8 drops of bromophenol blue indicator. If the solution is blue, pipet an additional 10.0 mL of 0.1 M HCl into the solution and boil again. Repeat as often as necessary. Record the total volume of HCl that is added to the antacid.

Obtain about 75 mL of a standardized 0.1 M NaOH solution. The solution may have been previously prepared by the stockroom personnel. If not, prepare a standardized 0.1 M NaOH solution, as described in Experiment 9. Consult with your laboratory instructor.

1. **Prepare the Buret for Titration.** Prepare a clean buret. Rinse the clean buret with two 3- to 5-mL portions of a standard NaOH solution. Record the actual molar concentration of the NaOH on the Report Sheet. Fill the buret with the NaOH solution; be sure no air bubbles are in the buret tip. Wait for 30 seconds, then read and record its initial volume (±0.02 mL).

2. **Titratre the Sample.** Once the antacid solution has cooled, titrate the sample with the NaOH solution to a blue endpoint. Watch closely, the endpoint may only take a few milliliters, depending on the concentration of the antacid in the sample. When a single drop (or half-drop) of NaOH solution changes the sample solution from yellow to blue, **stop**. Wait for 30 seconds and then read the final volume of NaOH solution in the buret.

3. **Repeat the Titration of the Same Sample.** Refill the buret and repeat the experiment.

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3If the sample is a tablet, swirl to dissolve. Some of the inert ingredients—fillers and binding agents used in the formulation of the antacid tablet—may not dissolve.

4Bromophenol blue is yellow in an acidic solution and blue in a basic solution.
4. Analyze Another Antacid. Perform the experiment, in duplicate, for another antacid. Record all data on the Report Sheet.

Disposal: Dispose of the solutions from the analysis in the sink, followed by a generous flow of water.

Cleanup: Discard the remaining NaOH titrant as directed by your instructor. Flush the buret several times with tap water and dispense through the buret tip, followed by several portions of deionized water. Dispose of all buret washings in the sink.
Date _____  Lab Sec. _____  Name _____________________________  Desk No. __________

1. Write the balanced equation for the reaction of one mole of the active ingredient in Rolaids with an excess of hydrogen ion. (Hint: Consider Equations 26.4 and 26.5.)

2. a. What is the color of the bromophenol blue indicator in an acidic solution?

b. What is its color in a basic solution?

c. Describe the color change that occurs at the endpoint in this experiment.

3. Why do some antacids cause gas to accumulate in the stomach? What is the gas?
4. A 25.00-mL volume of 0.104 M HCl is added to a sample of an unknown base. The HCl not neutralized (the excess HCl) by the base is titrated to a bromophenol blue endpoint with 10.70 mL of 0.0840 M NaOH. How many moles of unknown base (antacid) are present in the original sample?

5. The label of the liquid antacid Gaviscon reads that each tablespoonful (15 mL) contains 95 mg of aluminum hydroxide and 412 mg of magnesium carbonate. How many moles of hydronium ion can be neutralized by one tablespoonful of Gaviscon?

6. Each 5-mL teaspoonful of Extra Strength Maalox Plus contains 450 mg of magnesium hydroxide and 500 mg of aluminum hydroxide. How many moles of hydronium ion are neutralized by 1 teaspoonful of Extra Strength Maalox Plus?
# Experiment 26 Report Sheet

## Antacid Analysis

<table>
<thead>
<tr>
<th>Date</th>
<th>Lab Sec.</th>
<th>Name</th>
<th>Desk No.</th>
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### Commercial Antacid

<table>
<thead>
<tr>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Trial 1</th>
<th>Trial 2</th>
</tr>
</thead>
</table>

1. Mass of flask + antacid (g)  
2. Mass of flask (g)  
3. Mass of antacid (g)  
4. Volume of HCl added (mL)  
5. Molar concentration of HCl (mol/L)  
6. Molar concentration of NaOH (mol/L)  
7. Buret reading, final (mL)  
8. Buret reading, initial (mL)  
9. Volume of NaOH (mL)

### Calculations

<table>
<thead>
<tr>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Trial 1</th>
<th>Trial 2</th>
</tr>
</thead>
</table>

1. Amount of HCl added, total (mol)  
2. Amount of NaOH added (mol)  
3. Amount of base in antacid (mol)  
4. \( \frac{\text{mol base in antacid}}{\text{mass of antacid}} \) (mol/g)  
5. \( \frac{\text{average mol base in antacid}}{\text{mass of antacid}} \) (mol/g)  
6. Cost per gram of antacid (cents/g)  
7. antacid effectiveness \( \frac{\text{mol base in antacid}}{\text{cent}} \)  
8. Best buy
Laboratory Questions

Circle the questions that have been assigned.

1. If the CO₂ is not removed by boiling after the 0.1 M HCl is added, how does this affect the amount of NaOH required to reach the bromophenol blue endpoint? Explain. (Hint: CO₂ produces a low concentration of carbonic acid, H₂CO₃, when dissolved in water.)

2. If the antacid selected for analysis is known to contain only milk of magnesia, how could the Experimental Procedure be modified to expedite the analysis?

3. a. Write a balanced equation representing the antacid effect of sodium citrate, Na₃C₆H₅O₇. Citric acid is a weak triprotic acid, H₃C₆H₅O₇.
   b. Is sodium citrate more or less effective per mole than Mg(OH)₂? Per gram of Mg(OH)₂? Show calculations.

4. Assuming that stomach acid is 0.100 M HCl, how many grams of acid would be neutralized by an antacid tablet that contains 0.500 g of CaCO₃, the principal antacid in Tums.

*5. It is desired to increase the pH of the stomach acid from 1.0 to 2.0 with an antacid. Assuming stomach acid to be hydrochloric acid and the volume of the stomach to be 1.0 L, how many grams of CaCO₃ are necessary for the pH change?