Antacid Experiment Teacher Guide

Learning Objectives:

- Use of pH meters, pH indicators
- Use of burets and dosimeters to add reagents in titrations
- Use of potentiometric and colorimetric titrations to determine potency of a commercial Over The Counter (OTC) Antacid
- Calculation of the amount of calcium carbonate contained in an antacid tablet

Background:

The purpose of this experiment is to demonstrate some principles of analytical chemistry through the practical use of pH, pH meters, and pH indicator solutions. pH titrations (known as potentiometric titrations) can be useful tools in the determination of potency or the amount of material contained within a specific product such as an antacid. These same titrations can be carried out through the use of an indicator solution that changes color at a specific pH to determine the end point of the titration (colorimetric titration).

Antacids come in many forms, primarily consisting of some compound which neutralizes stomach acid. The most common compound found in antacids is calcium carbonate (CaCO₃). Eating too much or eating foods that may not agree with you can lead to overproduction of stomach acid (HCl) resulting in an upset stomach. You might decide to take an antacid to relieve that upset feeling.

In this experiment, a large drug store wants to market its own brand name antacid. It wants to sign a deal with a CMO (contract manufacturing organization) that will make the calcium carbonate antacid for them. Before they sign the multi-million dollar deal, they have requested production samples to be tested at an independent lab. They have selected our lab, East Hanover Analytical Labs (EHAL), to do this testing. We will determine the amount of active antacid product by titration, using pH indicators and pH meters to determine the potency of the antacid and then compare the results against the specifications set for this new product’s desired label claim.
Principles:

Basic compounds such as carbonates react with acids. As seen below in the chemical equation, calcium carbonate reacts with hydrochloric acid (HCl) to form carbon dioxide gas (CO₂), water (H₂O), and calcium chloride (CaCl₂).

The experiment performed is described by the following equation. The addition of the acid converts the calcium carbonate (pH 9-10 in water) to carbon dioxide and causes the pH of the solution to change. This change in pH can be monitored by either the use of an indicator solution or a pH meter.

Chemical Equations:

\[
\text{CaCO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{Ca(OH)}_2 + \text{CO}_2 \\
\text{Ca(OH)}_2 + 2\text{HCl} \rightleftharpoons \text{CaCl}_2 + 2\text{H}_2\text{O}
\]

Balanced equation: \[
\text{CaCO}_3 + 2\text{HCl} \rightleftharpoons \text{CaCl}_2 + \text{CO}_2 + 2\text{H}_2\text{O}
\]

And: \[
\text{CO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3 \text{ (carbonic acid)}
\]

In this experiment, we will learn to use pH indicators, pH meters, and titrametric equipment, such as a dispensing buret or a dosimeter, to perform the titration.

From the chemical equation above, it will take two (2) moles of hydrochloric acid (HCl) to neutralize 1 mole of calcium carbonate. We can use this information to determine the amount (in weight) of calcium carbonate neutralized by a volume of acid of known concentration.

Target pH

pH 7 is neutral pH, and so pH 7 would seem to be the target pH to use to determine the end point of the total conversion of calcium carbonate to calcium chloride.

From the equations above, we see that CO₂ and H₂O are produced in the destruction of CaCO₃. Carbon dioxide (CO₂) is a gas at STP (standard temperature of 25 °C and pressure of 1 atmosphere) but has some solubility in water. In water, it
forms carbonic acid, a very weak acid that has a pH of 3.9. Because of this our targeted end point of the titration of calcium carbonate is pH 3.9 while the titration mixture is fresh.

**Question:**

What do you think is the pH of seltzer or carbonated water? ___________

What about sodas like cola (Hint: phosphoric acid is added to soda. It is a stronger acid than carbonic acid)? ______________________

If left open, carbonated water becomes ‘flat’. Why does it become flat? What do you think happens to the pH of the flat carbonated water? ______________________

___________________________________________________________________

**pH Electrodes and Measurement of pH**

The measurement of pH (power or potential of hydronium) is based upon the measurement of the concentration of hydrogen ions in solution. The ion concentration is measured using a pH meter and pH electrode, which measures an electrical potential of the solution. The greater that the concentration of hydrogen ions (H\(^+\)) is, the greater the potential measured, and thus the lower the pH. This potential measurement is why the type of titration is known as a **potentiometric titration**.

The concentration of hydrogen ions is described in terms of the molar concentration in solution. These concentrations range from

\[ 1 \times 10^0 \text{M to } 1 \times 10^{-14} \text{M} \]

pH is defined by the equation:

\[ \text{pH} = - \log [H^+] \]

where [H\(^+\)] is the hydrogen ion concentration in moles/liter (M)

**Example:**  \[ \text{pH} = - \log [1 \times 10^{-7} \text{M}] = \text{pH 7 (neutral pH)} \]
So..., pH is measured in a range of values from 0 (very acidic) to 14 (very alkaline or basic)

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</table>

Acidic Neutral Basic

**Question**

- Predict the pH of the following items? (Acidic, Neutral, Basic)

Water

Vinegar

Ammonia

**Colorimetric Titrations**

Colorimetric titrations are again based around the change of pH of the solution and the formation of ions. However, in this case it is the indicator that changes form due to the pH of the solution. It is the formation of the ionic form that produces changes of color to the compounds used.

Indicators are useful in the titration of such things as strong acids, strong bases, weak acids, and weak bases. For today’s experiments, we will be titrating a weak base (carbonate) with a strong acid (HCl). If we plot the amount of acid added vs. the pH oh the solution I an acid-base titration, it should result in a titration curve that looks like the illustration below.
In this type of titration, you should select an indicator that changes color at a range at or near the equivalence point. In our case, this change should occur at pH 4.0.

Shown in the table below are some common and useful indicators and the pH range in which they would change color. Compare this table to the previous curve to see how they match up.

**Common Indicators used in Titrations**

<table>
<thead>
<tr>
<th>Indicator</th>
<th>pH 2</th>
<th>pH 3</th>
<th>pH 4</th>
<th>pH 5</th>
<th>pH 6</th>
<th>pH 7</th>
<th>pH 8</th>
<th>pH 9</th>
<th>pH 10</th>
<th>pH 11</th>
<th>pH 12</th>
<th>pH 13</th>
<th>pH 14</th>
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<tbody>
<tr>
<td>Methyl Red</td>
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<td>Methyl Orange</td>
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<td>Bromophenol Blue</td>
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<td>Bromothymol Blue</td>
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<td>Bromocresol Green</td>
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<tr>
<td>Phenolphthalein</td>
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<tr>
<td>Red Cabbage</td>
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</tbody>
</table>

Unfortunately, colorimetric titrations have some limitations, primarily due to interferences with other materials in solutions. For example, a significant number of particles in solutions may cloud the solution and make it difficult to see a color change.
Describe any other type of limitation that might cause colorimetric titrations not to be a useful tool for the determination of an Antacid?

_________________________________________________________________

_________________________________________________________________

Predict which method (pH or colorimetric) will be the best for analyzing the antacid tablets? What issues might you encounter?

_________________________________________________________________

_________________________________________________________________

Scenario:

A major box store wants to market its own labeled antacids. It has contracted with a generic manufacturer to prepare 750 mg tablets in a variety of colors and flavors and has requested production samples. Having received several bottles of the antacid, it has sent them to S2S Analytical Labs for independent testing.

Our job is to determine the strength (potency) of the tablets and determine if they are trustworthy manufacturers.

Each lab team will be given two tablets containing 750 mg of active ingredient (calcium carbonate). The team will then analyze the tablets for the amount of calcium carbonate (CaCO₃).

Note: The size and color of the tablet does not indicate the strength. Many times larger does not mean stronger, only more of other fillers may have been added.

Procedure:

1. Record the color of the antacid provided
2. Using a mortar and pestle, carefully grind a single tablet into a powder such that the powder is uniform and free of lumps
3. Transfer the powder to the container provided and using a spatula scrape any remaining powder from the mortar into the beaker.
Safety Note: This experiment will use 1M Hydrochloric Acid. Please use extreme caution when working with this solution. Ask for help.

4. Add approximately 40 mL of water to the beaker and a magnetic stir bar.

At this point, your team will either use the pH meter/dosimeter or the pH indicator solution/burette. Record your data on the worksheet provided.

5. Verify that the buret or dosimeter contains 1.0 M Hydrochloric Acid and record the initial volume on the Worksheet (1).

6. pH Meter – Place the pH electrode in your solution and titrate the antacid until the pH is stable below at pH 3.8-4.0. Record the final volume (2).

7. Indicator Solution – Add 10 drops of the indicator solution and titrate the antacid until you see a visual stable color change for the indicator used (green to yellow). Record the final volume on the buret (2).

8. Calculate the amount of calcium carbonate in your sample.

9. Repeat steps 1 – 8 on a new sample.

10. Record your calculated potencies on the excel spreadsheet on the lab computer.

**Titration Calculation Worksheet**

\[
\text{CaCO}_3 + 2\text{HCl} \rightleftharpoons \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}
\]

<table>
<thead>
<tr>
<th>Claimed Tablet Size = 750 mg CaCO(_3)</th>
<th>Tablet 1</th>
<th>Tablet 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Color of tablet</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 Initial Buret or Dosimeter reading (mL)</td>
<td>mL</td>
<td>mL</td>
</tr>
<tr>
<td>2 Final Buret or Dosimeter (mL)</td>
<td>mL</td>
<td>mL</td>
</tr>
<tr>
<td>3 (Line 2) – (Line 1) = Total Volume of HCl added</td>
<td>mL</td>
<td>mL</td>
</tr>
<tr>
<td>4 (Line 3) \times \text{Concentration of HCl (1.0M)}(^1) = \text{mmol of HCl used}</td>
<td>mmol</td>
<td>mmol</td>
</tr>
<tr>
<td>5 (Line 4) / 2(^2) = \text{mmol CaCO}_3</td>
<td>mmol</td>
<td>mmol</td>
</tr>
<tr>
<td>6 (Line 5) \times 100.1 = \text{Weight of CaCO}_3 (100.09)</td>
<td>mg</td>
<td>mg</td>
</tr>
</tbody>
</table>
**Calculation:**

1) The concentration of the HCl is 1.0 M/L = 1.0 mmol/mL

2) 1 mmol of carbonate is destroyed by 2 mmol of acid (see equation above)

Weight of calcium carbonate CaCO₃ (mg) = 
\[ \text{[volume of HCl titrant } 3 \text{ (mL)}] \times \text{[1.0 mmol/mL]}/[2 \text{ (mmol of HCl per mmol of CaCO₃)]} \times \text{[100.09 mg/mmol (MW)]} \]

Tracking Units: \[ \text{mg} = \frac{\text{(mL)}}{\text{(mL)}} \times \frac{\text{(mmol)}}{\text{(mmol)}} \]

Crossing out: \[ \text{mg} = \frac{\text{(mL)}}{\text{(mL)}} \times \frac{\text{(mmol)}}{\text{(mmol)}} \]
Instructor Notes

- Two students operate as a team. However, if numbers permit, high school students are encouraged to operate alone.
- The teams are spread evenly across the two modalities. If there are odd numbers of teams, put the odd number on the indicator/buret side (this side always seems to finish faster)
- The first student is taught the technique and how to do the calculations while the second student watches. The first student then becomes the ‘expert’ and helps the second student while the instructor watches.
- The students are told approximately how much acid ‘might’ be needed and shown the calculation as part of the introduction
- The students should practice the quick touch on the dosimeter or the rapid 180° rapid turn of the stopcock on the buret at the start of the addition of the acid
- The titration with the indicator or pH meter bounces from yellow to green or pH 2 to 4.5 because not all of the calcium carbonate dissolves in the volume of water used. The end point is when the pH remains at 3.8 to 4.0 or a persistent yellow remains for longer than 20 seconds.
- The teams are presented with one tablet at a time. This allows substitution of a smaller (500 mg) or larger tablet (1000 mg), which the students have yet to notice, that introduces a production error if desired. Some schools get tight titration numbers that you will notice with the first class, so they are good targets for the substitution.
- The indicator/buret method works only for yellow or white tablets. One orange tablet is given per class to only one team as a second tablet to teach them the limitations of the indicator method. They will not see the color change and thus over titrate and then they can explain their result to the class during discussion/questions.
- Accurate titrations are great but leave little to talk about during the review. As each session adds data, their data can be included in the discussion. School classes with a greater spread allow the summary and recommendation to include:
  - the need to greater sampling
  - talking of production problems
    - mixing of ingredients
    - making of tablets of uniform size)
  - the role of analytical labs in assuring safe and accurate production of any medicine
  - the role of titrations in assaying any pharmaceutical
  - the role of the FDA as the safety police
    - they carry badges and a congressional warrant
    - they cannot be refused or delayed admittance to review and investigate
they can review anything they want to review or interview anyone they want to interview at any time of their choosing
they can issue a recall, issue fines, or shut down an operation
they also review, on site, operations worldwide. If refused, they can quarantine materials made off-shore when they are imported

The final question to the students is whether they recommend continuation of the manufacturing contract. Why or why not? Can manufacturing recommendations be made based on the titration data? How could the brand name be harmed by poor manufacturing controls?

Demonstration of red cabbage juice indicator is best done when students are finished and waiting for others to finish.