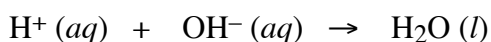


Experiment 7: Titration of an Antacid

Objective: In this experiment, you will standardize a solution of base using the analytical technique known as **titration**. Using this standardized solution, you will determine the acid neutralizing power of a commercially available antacid tablet.

Introduction

An understanding of the properties of acids and bases is an essential part of understanding chemical reactions. In aqueous solutions, a compound that produces H^+ ions upon dissolution is termed an **acid**. A compound that produces OH^- ions when dissolved in water is called a **base**. The reaction of an acid and base is a **neutralization** reaction, the products of which are a salt and water. In an aqueous solution, virtually all of the OH^- ions present will react with all of the H^+ ions which are present:



Because this reaction is essentially quantitative, it is possible to determine the concentration of an acid or base in an aqueous solution with high accuracy.

When a solution of hydrochloric acid, HCl, is *exactly* neutralized with a solution of sodium hydroxide, NaOH, the number of moles of NaOH used will equal the number of moles of HCl originally present. The following relationship then holds true:

$$\text{moles}_{\text{NaOH}} = \text{moles}_{\text{HCl}}$$

$$(M_{\text{NaOH}})(V_{\text{NaOH}} \text{ in liters}) = (M_{\text{HCl}})(V_{\text{HCl}} \text{ in liters}) \quad \text{Eq. 1}$$

where M = concentration in molarity and V = volume. If three of the above quantities are known, the fourth can be calculated.

In order to determine when a solution has been *exactly* neutralized, an **acid-base indicator** is used which changes color in a certain pH range (pH is a scale used to measure acidity). This color change is termed the **endpoint** of the titration. Because the pH of a neutral solution is 7, an indicator that changes color near this pH should be used for an acid-base titration. **Phenolphthalein** indicator changes color in the range pH = 8.3 – 10.0 and can be used to determine when the correct amount of base has been added to an acidic solution to exactly neutralize it.

Standardization of a Sodium Hydroxide Solution

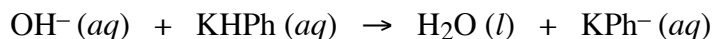
In order to determine the concentration of an acidic or basic solution, it is necessary to know the number of moles of acid or base that are required to neutralize it. This quantity can be calculated by accurately weighing a solid sample of an acid or a base, dissolving it in water and **titrating** this solution; that is, adding the solution of unknown concentration to it until the endpoint has been reached (see Tro, pp 163-165).

It is difficult to accurately weigh sodium hydroxide since it is **hygroscopic** (absorbs water readily from air). A solution of NaOH is usually *standardized* using an acid known as a **primary standard**. A primary standard must satisfy the four following criteria:

1. Solid compound that is not hygroscopic and can be easily handled

2. Is available in very pure form
3. Stable
4. Has a medium to high molecular weight

For this experiment, a solution of NaOH, which has an approximate concentration of 0.1 M, will be standardized using potassium acid phthalate, KHP. The molecular weight of KHP is 204.23 g/mole, and it has one **acidic proton**, which will react quantitatively with OH⁻:



For the highest accuracy, a sample size is chosen such that it will consume as large a volume of the base as possible without exceeding the capacity of the buret. If a 25 mL buret is used, the amount of KHP is chosen such that it will require approximately 20 mL of 0.1 M NaOH solution to reach the endpoint. Thus, about 0.002 moles, or **0.4 g**, of KHP is needed.

At the endpoint, the number of moles of NaOH equals the number of moles of KHP used:

$$M_{\text{NaOH}} = \frac{\text{moles KHP}}{V_{\text{NaOH}} \text{ in liters}} \quad \text{Eq. 2}$$

or

$$M_{\text{NaOH}} = \frac{\text{g, KHP}}{204.23 \text{ g/mole}} \times \frac{1000 \text{ mL/L}}{\text{mL, NaOH}} \quad \text{Eq. 3}$$

Once the NaOH solution has been standardized, it can be used to determine the acid neutralizing capacity of an antacid tablet.

Determination of the Acid Neutralizing Capacity of an Antacid Tablet

The stomach has an acidic interior generated by dilute HCl, "stomach acid", which insures proper digestion. When the acidity of the stomach becomes high enough to cause discomfort, brought about by the ingestion of certain types of food, an antacid preparation can be taken to neutralize the excess stomach acid. The active ingredient in every antacid is a base, the most common being metal hydroxides, metal carbonates or a mixture of the two. Table 1 lists the active ingredients in several commercial brands of antacid.

The acid neutralizing capacity of a tablet is the amount of hydrochloric acid that it can neutralize. It is the quantity which is referred to in some advertisements when it is stated that the tablet "neutralizes *x* times its weight in stomach acid". This capacity can be determined by a technique called **back-titration**. A known amount of antacid is dissolved in an excess of HCl, and then the excess acid is back-titrated with standardized NaOH solution. When the endpoint is reached, the number of moles of acid which was added to the antacid sample is equal to the number of moles of base present, NaOH *plus* the antacid. Therefore, the number of moles of HCl which was neutralized by the antacid is equal to the total number of moles of HCl added *minus* the number of moles which were neutralized by the NaOH:

moles acid neutralized = (moles of HCl added) – (moles of NaOH required for back-titration)

$$= (M_{\text{HCl}} \times V_{\text{HCl}}) - (M_{\text{NaOH}} \times V_{\text{NaOH}}) \quad \text{Eq. 4}$$

where M = molarity and V = volume in liters.

Table 1. Brands of antacid tablets and their major ingredients

Brand Name dose	Major ingredient	Recommended
Alka-Seltzer	NaHCO_3	1 or 2 tablets
Digel	$\text{Al}(\text{OH})_3$; $\text{Mg}(\text{OH})_2$	2 tablets
Phillips' Magnesia	$\text{Mg}(\text{OH})_2$	2-4 tablets
Maalox	$\text{Al}(\text{OH})_3$; $\text{Mg}(\text{OH})_2$	2 tablets
Rolaids	CaCO_3 ; $\text{Mg}(\text{OH})_2$	1-2 tablets
CVS brand	CaCO_3	2-4 tablets

Procedure

Water and NaOH solutions can absorb carbon dioxide, CO_2 , gas from the air, which will react with water to form carbonic acid, H_2CO_3 . Prepare approximately 250 mL of CO_2 -free water by boiling on a hot plate for 5 minutes. **CAUTION:** Set the boiling water well back from the bench edge to avoid accidents and burns. Allow the water to cool in a beaker, *covered* with a watchglass.

Study TECH sections II.E and F. Always handle the buret with care as it is quite expensive. Your buret must be scrupulously clean. A dirty buret can be identified by water drops adhering to the interior walls. If your buret is dirty, remove the stopcock, and using a brush, scrub the buret thoroughly from both ends with soap and water. Rinse with tap water followed by distilled water.

Pour approximately 150 mL of NaOH solution into a *dry* 250 mL beaker (avoid undue exposure of the base solution to air by keeping the storage bottle tightly capped). **Sodium hydroxide is corrosive!** The buret must be rinsed out with the solution being used before it is filled to prevent dilution of the solution. In this case, three *small*, successive portions of NaOH solution should be used to thoroughly rinse down the sides of the buret. The buret is filled to a point above the "0" mL mark with NaOH solution. In order to fill the tip of the buret with liquid, the solution is drained out of the bottom until the meniscus lies between the "0" and "1" mL marks. The initial buret reading can now be recorded **to the nearest 0.01 mL**. If you have any doubts as to your ability to read the buret correctly, ask your instructor to check your initial reading.

Standardization of NaOH solution

Accurately weigh out *three* samples of approximately 0.3-0.4 g of primary standard potassium hydrogen phthalate, KHP, which has been previously dried at 120°C . Do not use more than 0.4 g. To obtain an accurate mass, weigh the samples on weighing paper, slide them into three separate, clean (but not necessarily dry) 250 mL Erlenmeyer flasks and *reweigh* the paper to account for any KHP which may remain on it. Label the flasks with the corresponding mass for identification.

Dissolve the KHP sample with the smallest mass in about 50 mL of CO_2 -free water, and add 2-3 drops of 1% phenolphthalein indicator. Begin adding the approximately 0.1 M sodium hydroxide solution from the buret while **continuously swirling** the flask contents. Do not open the stopcock completely. As the endpoint nears, a pink color will appear at the point where the NaOH mixes with the flask contents. This color will disappear with subsequent swirling. Placing a white piece of paper under the flask will aid in observing the color change. When the color persists for 30 seconds after swirling, the endpoint has been reached. The color will fade after some time due to absorption of CO_2 from the air. If a deep pink color results, the endpoint has been overrun. Just prior to the endpoint, the flask walls should be rinsed down with a stream of distilled water from your wash bottle. It is possible to add "half-drops" of solution from the buret. Open the stopcock until a drop just forms at the tip of the buret. Touch the drop to the side of the flask and wash it down with distilled water.

When the endpoint is reached, record the final buret reading **to the nearest 0.01 mL**. Refill the buret so that you do not run out of NaOH solution in the middle of the next titration. Dissolve your second sample of KHP in about 50 mL of CO_2 -free water and repeat the titration procedure.

Since you titrated the smallest sample first, you know that the second and third titrations will require a larger volume of NaOH solution to reach the endpoint, and therefore these titrations may be done more rapidly. Most of the volume can be added quickly with constant swirling followed by the final drop by drop procedure used with the first sample. Your three determinations should not differ by more than 0.002 M (use equation 3 to do this calculation). In the case of poor precision, an additional sample may be run if there is time. Use an average of these molarities for analyzing the antacid in the next part of the experiment.

Back-titration of an antacid

Before you begin this part of the procedure, make a prediction about which tablet will be more cost effective and which tablet will be more weight effective. Write your predictions in your notebook.

Choose a brand and obtain 2 antacid tablets. Avoid touching them with your fingers as much as possible. Record the brand name, cost per package and number of tablets per package. Weigh each tablet separately on weighing paper to the nearest 0.001 g. Transfer each tablet to a 250 mL Erlenmeyer flask and label the flasks with the corresponding masses.

From the 50 mL dispensing buret, add approximately 12 mL of standardized HCl to a flask containing one of the antacid tablets. Record *precisely* how much acid was added using the initial and final readings of the buret. Also, be sure to record the *exact* molarity of the HCl solution written on the label of the bottle. **Hydrochloric acid is corrosive!**

Rinse down the inner walls of the flask with 40-50 mL of distilled water, and swirl the flask. You may need to crush the tablet with a glass stirring rod to allow the reaction to go to completion, but *be sure to rinse all the solution and particles off the rod before removing it*. The tablet may contain an insoluble binder and filler which will not dissolve; however, be certain that no large chunks or chips of the tablet remain.

Carbonate ions in the sample will react with the HCl and produce CO₂. The CO₂ can be driven off by heating the contents of the flask just to boiling:



Adjust the hot plate to maintain a gentle boil for *exactly* 5 minutes. Remove from the heat and allow the flask to cool until it is comfortable to hold.

Add 8 drops of phenolphthalein to the flask and swirl to mix; then rinse down the sides of the flask with your wash bottle. If the solution is pink at this point, more acid must be added from the buret, 1 mL at a time, until the color disappears.

Titrate the solution, as in the first part of the experiment, to the pale pink endpoint using the standardized NaOH solution. Refill your buret and repeat the entire procedure with the other antacid tablet.

When you have finished all the titrations, drain the NaOH from your buret into the sink, rinse it thoroughly with distilled water and return it to your instructor. Clean up your bench area and sink.

You will be asked to compare the weight and cost effectiveness of the two brands of antacid tablets. Before you leave the laboratory, find a student who tested the brand that you did not test and exchange data with him/her.

Calculations

1. Calculate the average molarity of the NaOH solution from your three standardization trials.
2. *Calculations to determine the effectiveness of the antacid*
 - Calculate the average number of moles of HCl neutralized per antacid tablet.
 - Calculate the number of moles of HCl neutralized per gram of antacid (*weight effectiveness*).
 - Calculate the average cost of the antacid tablet that would be needed to neutralize 1.00 mole of HCl (*cost effectiveness*).
3. The HCl concentration in a hyperacidic stomach is 0.03 M. The volume of liquid in the stomach is 300 mL. How many tablets of the antacid which you analyzed would have to be taken to bring the concentration of HCl in the stomach to a more normal level of 0.0003 M? Show your work.

Questions

1. Using data from a student who analyzed the other brand of antacid, evaluate and compare the two brands on the basis of weight and cost effectiveness. Do your results match your predictions?
2. Refer to Table 1 to identify the major basic ingredient contained in both of the antacid brands that were titrated. Write equations for the reactions of these active ingredients with acid, H^+ , (for both Phillips' Magnesia and CVS brand).
3. Name some factors, *besides* weight and cost effectiveness, that one might consider when choosing one brand of antacid over another.
4. List possible sources of error in this experiment and for each, indicate whether the effect would be an erroneously high or low value for the neutralization capacity of the antacid.