

## EXPERIMENT 6

### **Bases, Acids and Antacids**

#### **INTRODUCTION**

The terms acid and base are used quite loosely in chemistry and may have slightly different meanings under different circumstances. We will generally be using these terms to refer to chemical substances that dissolve in water. Keep in mind that an acid is a substance that increases the amount of  $\text{H}^+$  ion in water and a base is a substance that increases the amount of  $\text{OH}^-$  ion in water. It is important to realize that as the amount of  $\text{H}^+$  ion increases in water, the amount of  $\text{OH}^-$  ion decreases by a proportionate amount. The product of their molar concentration in pure water or dilute aqueous solutions will be constant (note - the square brackets [ ] around a chemical species represents the concentration of that species):

$$K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$$

The relative acidity or basicity of aqueous solutions is measured as pH (notice that the symbol used is lower case p followed by upper case H), which is defined as:

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

What this means is that for every change in the concentration of  $\text{H}^+$  by a factor of 10, the pH will change by a factor of 1. When  $[\text{H}^+]$  is  $10^{-2}$  M, the pH is 2 and when  $[\text{H}^+]$  is  $10^{-5}$  M, the pH is 5. When the pH is 12, this means  $[\text{H}^+]$  is  $10^{-12}$  M, which is a very small amount (1 trillionth molar or 1/1,000,000,000,000). In contrast, the  $[\text{OH}^-]$  will be relatively large at pH 12,  $10^{-2}$  M.

We encounter acids and bases frequently in our daily routines, although we don't always think about the acid or base properties of the myriad substances we come in contact with each day. We have discussed acid rain in the course and saw how it fits into the larger picture of our environment. Many of the foods we eat are acidic or alkaline (alkaline is another word for basic that can be used to avoid confusion with the word basic when it is used to mean "fundamental"). Acidic foods tend to taste sour (e.g., many fruits, especially citrus and tomatoes). Many foods that taste bitter are alkaline. But because we can perceive so many different tastes, the taste test isn't always a good indicator of whether substances are acids or bases.

In this laboratory exercise you will measure the acidity of vinegar and the relative

effectiveness of a few commercial antacid tablets. Antacids are a major commodity in the industrialized world, accounting for many millions of dollars of sales annually in the US alone. Antacids are, in fact, bases that are used to neutralize acid produced by the stomach after an overindulgence in food. The companies marketing these products often lead the consumer to believe "acid indigestion" is a common occurrence that no one should have to put up with.

The stomach juices are normally very acidic, with a pH in the range of 1 to 2. This acidic medium in the stomach has several important functions. The acid environment is essential for the enzymes in the stomach to break down proteins in foods, especially in meats. It also destroys many of the bacteria and other microorganisms that are ever present on our foods. Antacids should not be used casually and certainly should not be used on a continual basis. The components of all antacids can disrupt the normal balance of nutrients in the body when used over a prolonged period of time. Antacids should not be used soon after a meal, because the stomach secretes the acid to digest food. Neutralizing the stomach acid too soon may result in greater discomfort later.

The amount of acid or base in a substance is readily measured by chemical reaction of that substance with a known quantity of its counterpart, i.e., reacting an acid with a known quantity of base or vice versa. This type of measurement is best made by the technique known as titration. The titration method involves careful addition of a standard solution that reacts chemically with the substance you wish to quantitate. In this case, you are going to add the antacid to a known amount of acid in solution and determine how much acid is left over by titration with a standard base; sodium hydroxide in this case. See Appendix I for a description of the titration apparatus.

### **Materials Needed**

Commercial vinegar in original bottle, three brands of antacid tablets, phenolphthalein indicator solution, Congo Red dye indicator solution, 0.5 N HCl solution, 0.5 N NaOH solution, 50 mL buret, 250 mL and 125 mL Erlenmeyer flasks, 100 mL graduated cylinder, 50 mL graduated cylinder, 5 mL pipet.

### **PROCEDURE**

Part 1. The Acid Content of Vinegar.

Vinegar is a product of the aerobic fermentation of sugars, usually from fruit juices, such as apple (cider vinegar) or grapes (wine vinegar). Under anaerobic conditions (no oxygen) the fermentation produces alcohol,  $\text{CH}_3\text{CH}_2\text{OH}$  (hard cider or wine), but if oxygen is allowed in, the by-product of this fermentation by yeast is acetic acid,  $\text{CH}_3\text{COOH}$ . In this exercise you will determine how much acetic acid is in vinegar.

Clean three 125 mL Erlenmeyer flasks and add exactly 10.0 mL of vinegar (measure with the 5 mL pipet filled to the mark twice). Add about 20 mL of water and then add several drops of phenolphthalein indicator and note the color of the solution. Fill a 50 mL buret with 0.50 N NaOH (sodium hydroxide) solution and eliminate air bubbles from the stopcock and tip of the buret (see Titration Apparatus, Appendix I). Record the volume of NaOH solution in the buret before starting the titration. Titrate the vinegar solution in one of the Erlenmeyer flasks by adding the NaOH solution from the buret. Notice any color change in the flask as you add the base solution. The color should disappear as you stop adding the NaOH solution and swirl the flask. Keep adding the NaOH solution from the buret until the color change in the flask persists. Add the solution from the buret slowly as it takes longer and longer for the color change to disappear. Eventually the addition of just a few drops will result in the color change persisting. Do not add any more NaOH solution than is necessary to accomplish a persistent color change. Record the final volume of solution in the buret when you reach this **end point** in the titration, the point where the color of the dye persists in the flask.

**Do not add too much base to the antacid solution or your results will not be accurate.**

Repeat this procedure two more times using the vinegar solution in the other two Erlenmeyer flasks. Make sure there is enough NaOH solution in the buret before you start each titration and record the starting volume and the final volume in the buret for each titration. Calculate the molar concentration of acetic acid in the vinegar as follows:

$$\text{Molar Concentration of Acetic Acid} = \frac{(\text{x mL NaOH added})(0.50 \text{ mole/L})}{(10 \text{ mL vinegar used})}$$

Determine the average molar concentration of acetic acid in the three vinegar samples. To calculate the percent concentration of acetic acid in vinegar, multiply by the molecular weight of acetic acid (60 g/mole) and multiply by 0.1, since the molarity is based on 1.0 L and percent

concentration is based on g/100 mL (100 mL = 0.1 x 1.0 L).

$$\% \text{ Acetic Acid} = (\text{molar concentration})(60 \text{ g/mole})(0.1)$$

Record your results in your notebook.

### Part 2. Neutralizing Capacity of Antacids.

Record the "active" ingredients for each antacid tablet in your notebook. Crush each antacid tablet separately (so they will dissolve faster) and add each to a separate 250 mL Erlenmeyer flask. Add 70 mL of distilled water to each flask with the crushed tablet. Add exactly 30 mL of 0.5 N HCl (hydrochloric acid) to each. The concentration of the hydrochloric acid is 0.5 equivalent per liter, meaning there will be 0.5 mole of  $\text{H}^+$  ion per liter of that solution, or 15 milliequivalents of acid in the 30 mL of acid added to the antacid tablet in the water. Swirl it to get the tablet completely dissolved in the water. Take note of any gas that might be evolved and how quickly each antacid tablet dissolves and record this observation in your notebook.

After you have prepared the antacid solution, refill the buret with 0.5 N NaOH solution and record the initial volume of base solution in the buret. Make sure the level of solution in the buret is in the range of graduation marks so you can record the initial volume.

Add several (5-10) drops of Congo Red indicator dye to the antacid solution and take note of the color. Titrate the acid-antacid solution by carefully adding base from the buret. Swirl the solution while adding the base, and turn off the stopcock when the color in the flask begins to turn red or pink. Go slowly when you begin to see a color change. Add a few drops of base from the buret, turn off the stopcock, swirl the flask and see if the color persists. When the pink color persists, you have reached the endpoint of your titration. Make sure the color persists for at least a minute. Record the final volume of base in the buret and determine the volume of standardized sodium hydroxide solution you needed for the titration.

**Note: Some of the antacids contain artificial colors that may cause the final color to be other than pink, but look for a change from a bluish hue to a reddish hue (e.g., greenish to orangish).**

The amount of acid neutralized by the antacid tablet can now be calculated. Since the

acid that you added is only partially neutralized by the antacid tablet, the NaOH is used to neutralize the remaining (or excess) acid. The amount of acid added is known (15 milliequivalents). The amount of base used to titrate the excess acid can be calculated and is equivalent to the amount of excess acid present. This is equal to the volume of base multiplied by its concentration, which is 0.5 equivalents per liter or 0.5 milliequivalents per milliliter. The milliequivalents of base in the antacid tablet will be the difference between the total acid added and the excess acid titrated with sodium hydroxide, or:

$$\text{meq antacid} = (15 \text{ meq acid}) - (0.5 \text{ meq NaOH per mL}) (\text{mL used})$$

Repeat the procedure for each antacid tested and plot the results as a bar graph. The bar graph should have meq of antacid on one axis and the brand of tablet on the other. Draw conclusions in your notebook about the relative strength of each brand of antacid tablet and how this is related to its chemical composition or "active" ingredients. Be sure to record in your notebook the composition of each antacid tablet with respect to its active ingredients.

Questions to answer in your notebook:

1. What is the purpose of the phenolphthalein indicator in the vinegar solution?
2. What is the role of Congo Red dye in the titration of antacid tablets?
3. From this experiment, which antacid tablet would you predict gives the most relief from excess stomach acid?
- 4.\* After comparing the amount of active ingredients in each tablet, which ingredients seem to give the most acid neutralization? Write the chemical formula for the active ingredients in each antacid tablet and show how many  $\text{H}^+$  ions can displace the cations in each formula by writing a balanced chemical equation for each reaction.