Acids, Bases and Antacids

Experiment #8

Objective: To be able to quantify the amount of acid in vinegar and a citric acid solution, and the equivalents of weak base in antacid tablets using the technique known as titration.

Introduction

We will use the Brønsted-Lowry definition for acids and bases in water. An acid is a substance that increases the amount of H^+ ion in water and a base is a substance that increases the amount of OH^- ion in water. As the amount of H^+ ion increases in water, the amount of OH^- ion decreases by a proportionate amount. The product of their molar concentration in pure water or dilute aqueous solutions will be constant:

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

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:

$$pH = -\log [H^+]$$

What this means is that for every change in the concentration of H^+ by a factor of 10, the pH will change by a factor of 1. When $[H^+]$ is 10^{-5} M, the pH is 5 and when $[H^+]$ is 10^{-2} M, the pH is 2. When the pH is 12, this means $[H^+]$ is 10^{-12} M, or 0.000000000001 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. Many of the foods we eat are acidic or alkaline (alkaline is another word for basic). Acidic foods tend to taste sour (e.g., many fruits, especially citrus fruits and tomatoes). Many foods that taste bitter are alkaline.

Vinegar is a product of the aerobic fermentation of sugars in fruit juices by yeast, such as apple (cider vinegar) or grape (wine vinegar). Under anaerobic conditions (no oxygen) the fermentation produces ethyl alcohol, CH_3CH_2OH (hard cider or wine), a more reduced product than sugar. In the presence of air (or oxygen), the yeast produce acetic acid, CH_3COOH, a more oxidized product than sugar. In this exercise you will determine how much acetic acid is in vinegar.

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. Some individuals experience a condition known as acid reflux, where stomach acid enters the esophagus and irritates the lining of the esophagus, causing what is commonly known as heartburn. Antacids are commonly consumed to neutralize the stomach acids and help to alleviate heartburn. 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 active ingredients in several antacids are shown in Table 8.1.

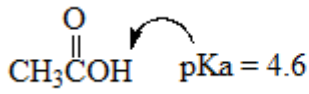
Table 8.1. Some common antacids and their active ingredients.

Antacid Brand Name	Active Ingredients (Weak Bases)	Equivalents/mole for active ingredient
Alka Seltzer	$NaHCO_3$	1
Maalox, Mylanta	$Al(OH)_3$ and $Mg(OH)_2$	3 2
Tums, Alka-2	$CaCO_3$	2
Rolaids	$AlNa(OH)_2CO_3$	4
Milk of Magnesia	$Mg(OH)_2$	2

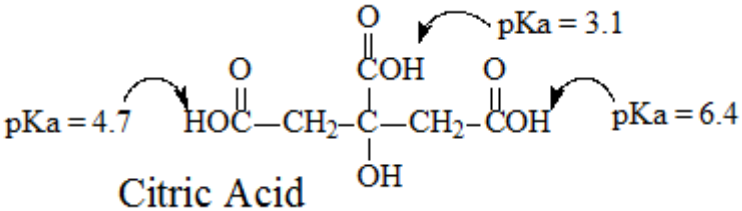
The amount of acid or base in a substance is readily measured by titration of that substance with a known quantity of its counterpart, *i.e.*, titrating an acid with a known quantity of base or vice versa. Concentrations of acids and bases are usually measured in equivalents per liter (eq/L) because some acids and bases have more than one acid group or base group per molecule (or eq/mole). Equivalents per mole are given for the antacids in the table above. During titration the end point is when equivalents of acid equals equivalents of base, *i.e.* neutrality occurs.

In this experiment you will determine the concentration of acetic acid in vinegar and also measure the buffering capacity or number of acid **equivalents** for a weak acid that is often used in foods and beverages, citric acid. Notice in the figure below that citric acid has 3 equivalents per mole, or 3 acidic H that can dissociate as H^+ ions. The alcoholic -OH group on the central C atom is not acidic. Acetic acid is also shown, which has 1 equivalent per mole. You will also add an antacid to a known amount of hydrochloric 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 and how to read a buret.

Another term that is used is pK_a for weak acids, where the pK_a is the pH where the acid group is half dissociated. You can think of the pK_a as the pH where a particular H^+ ion is removed from a weak acid. The pK_a for acetic acid is 4.6, whereas citric acid has three pK_a values of 3.1, 4.7 and 6.4, meaning it loses its first H^+ ion at pH 3.1, its second at pH 4.7 and its third H^+ ion at pH 6.4.



Acetic Acid



Citric Acid

Materials Needed

Commercial vinegar, three brands of antacid tablets, phenolphthalein indicator solution, Congo Red indicator solution, 0.50 M hydrochloric acid solution, 0.50 M sodium hydroxide solution, 0.20 M citric acid solution, 50 mL buret and buret clamps.

Procedure

A. The acid content of vinegar

1. Clean two 250 or 300 mL Erlenmeyer flasks and add exactly 5.0 mL of vinegar (measure with the 10 mL graduated cylinder) to each.

2. Add about 20 mL of deionized water to each flask and then add 5 drops of phenolphthalein indicator to each flask and note whether there is any color in the solution.

3. Add about 80 mL of 0.50 M sodium hydroxide (NaOH) solution to a clean 150 mL beaker. You will use this to fill the buret for all parts of this experiment.

4. Using a buret clamp that is supplied, mount a 50 mL buret on the bars at your bench and fill it with 0.50 M NaOH (sodium hydroxide) solution. Eliminate air bubbles from the stopcock and tip of the buret (see Appendix I) and bring the level of liquid to the zero mark or below zero so you can get an accurate reading of the volume. Remember that zero is on top and 50 mL is at the bottom of the buret, so you will be reading the volume from top to bottom. Record the level of NaOH solution in the buret to the nearest 0.1 mL on the report sheet before starting the first titration. If the level of liquid is at zero on top, you should record 0 and **not** 50. If the level is midway between 1 and 2 the volume would be 1.5 (**not** 2.5 and **not** 48.5)

5. 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. 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 to disappear. Eventually the addition of just a few drops will result in the pink color persisting. Do not add any more NaOH solution than is necessary for the pink color to remain. This is the **end point** for neutralization.

6. Record the final level 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 and determine the volume used. (You should have used between 6 and 12 mL of base from the buret).

7. Check to make sure you have enough NaOH solution in the buret before you begin the next titration. You know how much was used for the first titration and the next one should use about the same amount to reach the end point. If there is not enough, refill the buret with NaOH solution before starting the next titration.

8. Record the starting level for the second titration on the report sheet, which will be the final level for the first titration if you did not add more solution to the buret.

9. Titrate the second flask containing vinegar to the end point and record the final level in the buret for the second titration and determine the volume used.

10. Calculate the the volume of NaOH solution used for both titrations and calculate the molar concentration of acetic acid in the vinegar as follows:

(Remember equivalents of acid = equivalents of base at the end point).

$$\text{Molar Concentration of Acetic Acid} = \frac{(\text{x mL NaOH added})(0.50\ \text{mole/L})}{(\text{5 mL vinegar used})}$$

11. Determine the average molar concentration (mole/L) of acetic acid in the two vinegar samples.

12. To calculate the percent concentration of acetic acid in vinegar, multiply molar concentration by the molecular weight of acetic acid (60 g/mole) and by 0.10 L, since the molarity is based on 1.0 L and percent concentration is based on g/mL x 100%. (or g/100 mL = g/0.10 L).

    % Acetic Acid = (mole/L)(60 g/mole)(0.10 L)

13. Record your results on the report sheet and answer the questions for this part.

    You can discard the solutions in each flask in the sink and rinse the flasks well with water and use them for the next part.

B. Acid equivalents in citric acid

1. Using the 10 mL graduated cylinder, add exactly 5.0 mL of 0.20 M citric acid solution to each of two 250 or 300 mL Erlenmeyer flasks.

2. Add about 20 mL of deionized water to each flask and add several drops of phenolphthalein indicator to each flask.

3. Refill the buret with 0.50 M sodium hydroxide solution if necessary and record the initial level in the buret to the nearest 0.1 mL on the report sheet for the first titration of this part.

4. Titrate the first flask of citric acid solution and phenolphthalein indicator in it to the end point, where the color of the phenolphthalein solution changes to pink and just enough NaOH solution is added to get the pink color to persist.

5. Record the final level in the buret on the report sheet and calculate the number of mL of 0.50 M NaOH added and the number of milliequivalents of base needed to neutralize the citric acid.

6. Calculate the number of millimoles (mmol) of citric acid in each flask and record it on the report sheet for this part. (Since the citric acid solution is 0.20 mol/L x 5.0 mL = 1.0 mmol citric acid)

7. Calculate the number of equivalents of acid per mole of citric acid on the report sheet.

8. Record the initial level of NaOH in the buret for the second titration of citric acid, which should be the same as the final level for the first titration, unless you add more base to the buret. Make sure there is enough NaOH solution in the buret to complete this titration, since you know about how much it should require from the first titration.

9. Titrate the second flask containing 5 mL of 0.20 M citric acid solution, water and phenolphthalein indicator. Add just enough base to have the color of the indicator remain pink for at least one minute, record the final level of base in the buret and determine the volume of NaOH solution used.

10. Calculate the number of equivalents of acid per mole of citric acid for this titration on the report sheet.

11. Use a piece of universal pH paper to estimate the pH of each solution after you have reached the endpoint with phenolphthalein indicator and enter the pH on the report sheet.

12. You can discard the contents of the flasks into the sink and rinse the flasks 2 or 3 times with water for the next part.

C. Neutralizing Capacity of Antacids

Note: We will use equivalents or milliequivalents of acid or base in these calculations because some of the antacids can neutralize two equivalents (or more) of acid. One mole of H^+ ion is the same as one equivalent of H+ ion, whereas one mole of $Mg(OH)_2$ contains two equivalents of base (OH^-).

1. Obtain two antacid tablets of one brand and record the "active" ingredients for that antacid on the report sheet.

2. Crush each antacid tablet **separately** (so they will dissolve faster) and add each separate powder to separate 250 mL Erlenmeyer flasks. Add 50 mL of deionized water to each flask with the crushed tablet. Using the 50 mL graduated cylinder, add exactly 30 mL of 0.50 M HCl (hydrochloric acid) to each flask. The HCl solution contains 0.50 equivalents/L (eq/L).

3. The number of milliequivalents of H^+ ion in 30 mL of hydrochloric acid solution is given in the appropriate place on the report sheet. [(30 mL)(0.50 eq/L) = 15 meq; Note: mL x eq/L = meq]

4. Swirl the flask to get the powder completely dissolved in the water. Take note of any gas that might be evolved and how quickly each antacid tablet dissolves. Add 5 drops of Congo Red indicator dye to each antacid solution and take note of the color.

5. After you have prepared the antacid solutions, refill the buret with 0.50 M NaOH solution and record the initial level in the buret on the report sheet for the first antacid tablet. Make sure the level of solution in the buret is in the range of graduation marks so you can record the initial level to the nearest 0.1 mL.

6. 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 reddish color persists, you

have reached the endpoint of your titration. Make sure the color persists for at least a minute.

Note: Some of the antacids contain artificial colors that may cause the final color to be other than red or pink, but look for a change from a bluish hue to a reddish hue.

7. Record the final level of base in the buret and determine the volume of 0.50 M sodium hydroxide solution you needed for the first titration and calculate the number of milliequivalents of base added in this titration.

8. Calculate the amount of acid neutralized by the antacid tablet on the report sheet. Since the hydrochloric 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 milli-equivalents). 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 0.50 equivalents per liter or 0.50 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:

$$\text{meq antacid} = (15 \text{ meq acid}) - (0.50 \text{ meq NaOH per mL})\ (\text{mL used})$$

9. Repeat the procedure for the second antacid tablet. Make sure you have enough solution in the buret and record the level of NaOH solution in the buret before you start titrating. It should be the same as the final level for the first antacid titration unless you add more base to the buret. Record the data on the report sheet and calculate the average number of milliequivalents of base in the antacid tablets and answer the questions for this part.

10. You can dispose of the contents of each flask in the sink and rinse the flasks. Return any 0.50 M sodium hydroxide solution remaining in the buret and beaker to the stock bottle and rinse the buret at least 3 times with water, making sure to open the stopcock at the bottom of the buret with water in it to rinse the sodium hydroxide solution from the stopcock and tip of the buret.

    Return the clean buret and buret clamp to where you got them.

Name _____ Section _____

Acids, Bases and Antacids
Experiment #8 **Prelab Exercise**

1. Give a brief description of the difference between a strong acid and a weak acid, giving at least one example of a strong acid and one example of a weak acid (give specific chemical compounds for each).

2. Describe the difference between a strong base and a weak base and give at least one specific example of each (one strong base and one weak base).

3. Explain why an indicator dye is used in the titration of an acid with a base, or in the titration of a base with an acid. What is the function of the indicator dye? Give at least one example of an indicator dye used for acid-base titrations.

Name _____ Section _____

4. If 22.8 mL of 0.50 M sodium hydroxide is used to completely neutralize 10.0 mL of an unknown acid in a titration experiment, calculate the following:

a) How many millimoles of sodium hydroxide were added to the acid in this titration? Show your work to get this answer.

b) How many milliequivalents of sodium hydroxide were added to the acid in this titration?

c) How many milliequivalents of acid were present in the 10.0 mL of unknown acid that were titrated? [Note: equivalents of acid = equivalents of base at neutrality]

d) If the unknown acid has two equivalents per mole (2 eq/mol), what is the molar concentration of the unknown acid that was titrated?

Name _____ Section _____

Acid, Bases and Antacids
Experiment #8 **Data & Report Sheet**

A. Acid content of vinegar

Enter the buret readings for initial and final volume of sodium hydroxide solution in the buret for each titration and calculate the concentration of acetic acid in vinegar.

	Titration 1	Titration 2
(a) Initial volume		
(b) Final volume		
(c) Volume of NaOH used [b – a]		
(d) millimoles of NaOH used (0.50 mmole/mL)(mL used)		
(e) mmole of acetic acid neutralized		
(f) Volume of vinegar used	5.0 mL	5.0 mL
Concentration of acetic acid in vinegar (mmoles/mL = mol/L or M) [(e)/(f)]		
Average of 2 trials (M)		
% Concentration of acetic acid in vinegar [see formula step A-14]		%

A-1. From the structure of acetic acid given in the introduction, how many equivalents of acid are there per mole of acetic acid?

A-2. Vinegar is generally considered to be about 5% acetic acid by weight. How do your results compare with the expected value? If your value is much different from the expected value, check your calculations?

Name _____ Section _____

B. Acid equivalents in citric acid

Enter the buret readings for initial and final level of sodium hydroxide solution in the buret for each titration. Calculate the number of milliequivalents (meq) of NaOH used to neutralize the citric acid in each flask.

	Trial 1	Trial 2
(a) Initial volume		
(b) Final volume		
(c) Volume of NaOH used [(b) – (a)]		
(d) meq of NaOH used [(0.50 meq/mL)(mL used)]		
(e) Volume of 0.20 M citric acid added	5.0 mL	5.0 mL
(f) Millimoles of citric acid added [(0.20 mol/L)(5.0 mL)]		
(g) meq/mmol of citric acid [(d)/(f)]		
(h) Average meq/mmol for citric acid		
(i) pH determined with universal pH paper		

B-1. From your data how many protons (H^+ ions) can be donated by each molecule of citric acid when it reacts with a base or proton acceptor? (See (h) in the table for meq/mmol for citric acid).

B-2. What may account for any significant difference between the expected number of protons that can be released (see the structure of citric acid in the introduction) and the number that you determined in the titration? [Look at the structure of citric acid in the introduction and see how many acid groups are in the molecule. Did you stop adding base when there was just enough base to neutralize the citric acid, or did you add too much?]

Name _____ Section _____

C. Neutralizing Capacity of Antacids

What is the brand name and active ingredient(s) for the antacid you have chosen?

Brand name: Active ingredient(s):

Enter the buret readings for initial and final level of sodium hydroxide solution in the buret for each titration. Calculate the number of milliequivalents (meq) of NaOH used to neutralize the excess hydrochloric acid added to the antacid tablet in each flask and the meq of antacid in each tablet.

	Trial 1	Trial 2
(a) Initial volume		
(b) Final volume		
(c) Volume of NaOH used [b - a]		
(d) meq of NaOH used [(0.50 meq/mL)(mL used)]		
(e) meq of HCl added to antacid	15 meq	15 meq
(f) meq of acid neutralized by antacid tablet [(e) – (d)]		
Average meq of base in antacid tablets		

C-1. If hydrochloric acid is the main component of stomach acid and the pH of the stomach is 2.0, what is the concentration of HCl in stomach acid? [See introduction regarding definition of pH].

C-2. From the average number of meq of base in your antacid tablets, calculate the volume (in mL) of stomach acid that can be neutralized by one antacid tablet. HCl has one equivalent per mole and you just calculated the number of moles of HCl per liter in stomach acid, so how many liters (or mL) of acid are neutralized by the meq of base in the antacid? [Note: if you divide the number of meq of acid neutralized (line f in the table above) by the number of meq/L for the stomach acid (question B-1 above), you will get the number of liters (L)].