

## Learning Objectives

The objectives of this experiment are to...

- introduce the nature of acids and bases.
- introduce acid-base indicators, e.g., litmus, wide range indicator papers and the specific titration indicators of methyl orange, bromothymol blue and phenolphthalein.
- Introduce various acid/base reactions
- introduce titration as a means of determining the amount of an acid or base present.

## Background

A discussion is presented on the origin of acid base theory as originally defined by Svante Ahrennius and later defined by Bronsted and Lowry.

## Indicators

Indicators are defined as weak organic compounds in which the acid form has a different color than the base form and their use in determining what substances are acids or bases, and their use in titrations is discussed. Experiments are carried out to determine the acid range, transition range and base range of each indicator using buffer solutions from pH 2 to 12.

## Titration

The process of titration is defined and the concepts of titrant, analyte, end point and equivalence point are defined. pH is defined as  $\text{pH} = -\log [\text{H}^+]$  and the use of the *MicroLAB* interface and computer with a pH probe to measure the pH is discussed. A titration of acetic acid is carried out.

## Acid-Base Experiments

A series of experiments is carried out in well plates which explores the interaction of different acids and bases with different common substances.

## Data Analysis

Guidance is given in each part of the experiment to help the students obtain the goals of the lab.

## Instructor Resources Provided

- Sample Report Sheets providing the format to organize the data collection with sample data.
- Questions to consider, answer and turn-in with suggested answers.
- Tips and Traps section to assist the instructor with potential problems and solutions.
- Sample *MicroLAB* screen shots and graphs.
- Laboratory preparation per student station.

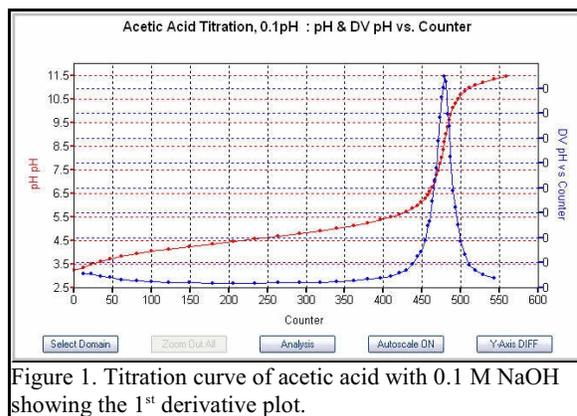


Figure 1. Titration curve of acetic acid with 0.1 M NaOH showing the 1<sup>st</sup> derivative plot.

## Experiment

# INTRODUCTION TO ACIDS, BASES AND TITRATION

## The CCLI Initiative

Computers in chemistry Laboratory Instruction

### LEARNING OBJECTIVES

The objectives of this experiment are to . . .

- introduce the nature of acids and bases.
- introduce acid-base indicators, e.g., litmus, wide range indicator papers and specific titration indicators.
- Introduce various acid/base reactions
- introduce titration as a means of determining the amount of an acid or base present.

### BACKGROUND

An acid was originally defined by Svante Ahrennius as a substance which furnishes hydrogen ion in solution, and a base as a substance which furnishes hydroxide ion in solution. Bronsted and Lowry further defined an acid as a proton donor and a base as a proton acceptor.

#### Indicators

We determine whether a substance is an acid or base by using "indicators." An indicator is an organic compound which can either accept or donate a proton, depending upon the hydrogen ion concentration of the solution it is placed in. Indicators have different colors depending on the hydrogen ion concentration. For example, litmus, a naturally occurring dye with this capability, is red in acid and blue in base. Other indicators have other colors. In addition, different indicators will change color at different hydrogen ion concentrations, or acidities. We will also utilize buffer solutions, which are solutions which have a specific acidity and tend to resist a change in that acidity.

#### Titration

Titration is a process in which a solution of known concentration is quantitatively added to a known volume of a solution of unknown concentration in order to determine its concentration. It requires the use of a buret, which is a glass tube of uniform bore with accurate calibrations capable of measuring any dispensed volume to a precision of 0.02 ml. The known volume of solution of unknown concentration to be titrated is known as the "analyte," and the solution in the buret of known concentration is termed the "titrant." The titration takes place by incrementally adding the titrant to the analyte solution.

As the titration proceeds, it is necessary to have some mechanism to identify when the proper amount of titrant has been added. Stoichiometrically, this occurs when chemically equivalent amounts of titrant have been added to the analyte. Macroscopically, this point is not generally observable, and it is necessary to measure it by some artificial means. In an acid-base titration such as we are doing here, there is a change in the acidity of the solution, and monitoring this change can serve the purpose. In practice, several methods

are available for determining this change in acidity. One is the use of indicators, as discussed above. The point in the titration at which the indicator changes from the acid color to the base color (titrating with a base) is defined as the "end point," and signals the end of the titration.

Another method of determining when the correct amount of titrant has been added is the use of a pH meter. pH is defined as the negative logarithm of the hydrogen ion concentration,

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

and ranges from 1 for very acid solutions ( $[\text{H}^+] = 10^{-1}$ ), to 14 for very basic solutions

( $[\text{H}^+] = 10^{-14}$ ). A pH meter measures the pH of the solution as the titration proceeds. We will use the **MicroLAB** interface and computer with a pH probe to measure the pH, and the system will graph the progress of the titration on the screen as the titrant is added

In this experiment, our titrant is sodium hydroxide, a strong base. We will be titrating acetic acid, a weak acid that only ionizes to about 2% in solution. We will then examine the titration curve as to its general shape, the initial and final pH values, and the pH at the equivalence point.

We will first explore the interaction of different acid concentrations with different indicators, then we will explore the interaction of different acids with different substances, then we will investigate the nature of the change in acidity of an acid by the incremental addition of a base.

## SAFETY PRECAUTIONS

### SAFETY GOGGLES MUST BE WORN AT ALL TIMES DURING THIS EXPERIMENT

- Hydrochloric Acid solution:** Toxic by ingestion and inhalation, strong irritant to eyes and skin.
- Nitric Acid solution:** Very dangerous to eyes and skin. Severe tissue damage upon ingestion.
- Sulfuric Acid solution:** Severely corrosive to eyes, skin and other tissue. Toxic, strong skin irritant. Powerful dehydrator causing blistering of the skin.
- Acetic Acid solution:** Corrosive liquid, skin burns are possible, very dangerous to eyes.
- Sodium Hydroxide solution:** Corrosive liquid, skin burns are possible, very dangerous to eyes.

The other chemicals are innocuous; however you should keep all chemicals away from eyes and mouth, wash hands after use and before leaving the laboratory, and use prudent laboratory practices at all times.

### BEFORE PERFORMING THIS EXPERIMENT . . .

...you will need a **MicroLAB** program capable of timing the collection of 5 ml titrant at about 2 - 4 seconds per drop. Your program should also be designed to collect pH data in relation to time and load this information into a spreadsheet. Use the **pH.Time.exp** selected from the **Titrations** tab.

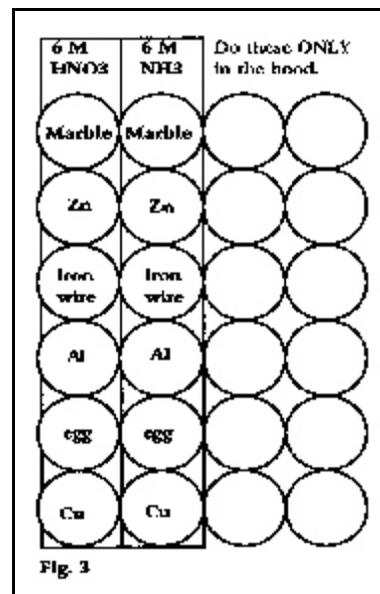
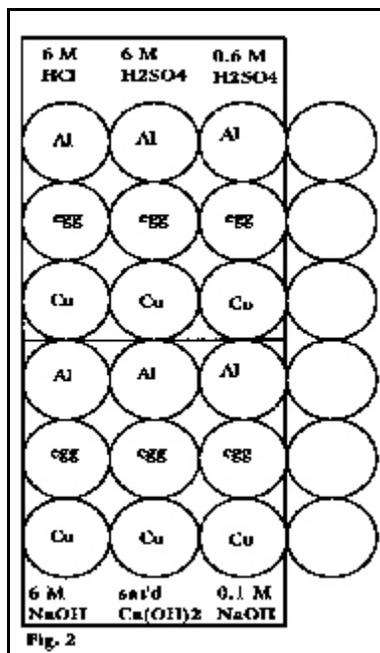
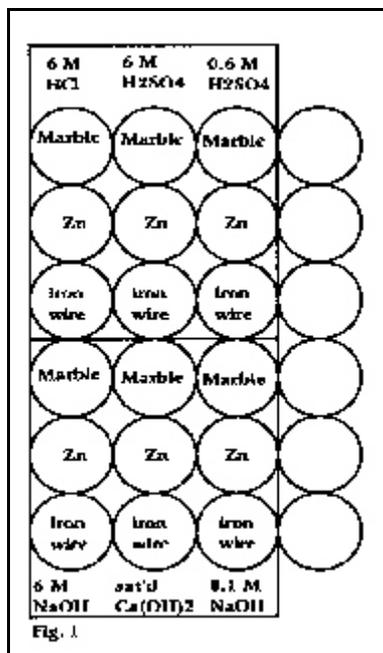
## EXPERIMENTAL PROCEDURE

### Acidity and Indicators: Determining the pH Acid, Transition and Base Range of Indicators

1. Add about 1 ml (1 plastic dropper full) of each buffer solution, 2 through 12, in separate wells of the well plate.
2. Dip a stirring rod into each buffer and touch it to a short piece of Alkacid paper, then match the color of the Alkacid paper against the chart on the side of the tube.
3. Under each buffer value on the report sheet, indicate whether the color you observed on the Alkacid paper lighter or darker than the color on the chart.
4. Add one drop of methyl orange indicator to each well, stir with a clean toothpick and note the pH **Acid**, **Transition** and **Base Range** over which the indicator makes its color transition. Indicate this on the buffer chart on the report sheet.
5. Empty the well plate into a large beaker, rinse and wipe dry, and refill with one ml each of buffer.
6. Repeat step four using bromothymol blue, then, then phenolphthalein, noting the **Acid**, **Transition** and **Base Range** for each on the report sheet.

### Acid/base reactions (using a 24-well plate)

1. Arrange six (6) small pieces each of marble chip, zinc, and iron wire in the 24 well plate as indicated in Figure 1.
2. Add 1 dropper of 6 M HCL, 1 dropper of 6 M H<sub>2</sub>SO<sub>4</sub>, and 1 dropper of 0.6 M H<sub>2</sub>SO<sub>4</sub> to each of the first three samples of marble, zinc and iron wire as indicated in Figure 1.
3. Observe the reaction of each substance with each of the acid solutions and record your observations in the report sheet table. Pay particular attention to differences in the reactions and to differences in reaction rates.
4. Hold a lighted match close to the surface of each of the liquids that show any bubbling in the wells and observe closely any change in the nature of the flame.
5. Add 1 dropper of 6 M NaOH, 1 dropper of sat'd Ca(OH)<sub>2</sub>, and 1 dropper of 0.1 M Na(OH) to each of the last three samples of marble, zinc and iron wire as indicated in Figure 1.



6. Again, observe the reaction of each substance with each of the base solutions and record your observations in the report sheet table. Pay particular attention to differences in the reactions and to differences in reaction rates.
7. Again, hold a lighted match close to the surface of each of the liquids that show any bubbling in the wells and observe closely any change in the nature of the flame.
8. Empty the contents of the wells into the discard beaker and rinse the cells out well with tap water, then a light rinse of distilled water and wipe dry with Kim Wipes.
9. Now arrange six (6) small pieces each of aluminum, six (6) sets each of a few drops of egg albumin and six (6) small pieces of copper in the 24 well plate as indicated in Figure 2.
10. Repeat steps two through eight above and record your observations on the report sheet.
11. Place two (2) pieces each of marble chips, zinc, iron wire, aluminum, 2 sets of a few drops of egg albumin and copper in separate wells as indicated in Figure 3. Move the well plate to a fume hood and add 1 ml of 6 M HNO<sub>3</sub> and 1 ml of 6 M NH<sub>4</sub>OH to each well and record your observations on the report sheet. Pay particular attention to differences in the reactions and to differences in reaction rates.

**Caution:** Do not put your face close to the well plate because a poisonous gas may be produced.

12. **Do not remove your well plate from the fume hood** but dump the solutions into the waste beaker in the hood labeled HNO<sub>3</sub> waste, rinse the well plate into the beaker with your wash bottle, then rinse in the sink and wipe dry.

## Titration

1. Check the buret for cleanliness and proper drainage as directed by your instructor.
2. Rinse the buret three times with about 5 ml amounts of the standard NaOH solution, tipping and rotating the buret to cover all of the inside surface of the buret with each rinse. Be sure to have a stopper in the end of the buret so you don't spill NaOH on the floor or counter top.
3. Fill the buret above the 0 mark with the NaOH solution.
4. Measure exactly 25 ml portions of the 0.1 M HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> solution (the analyte solution) in a graduated cylinder and pour into each of two or three 250 ml beakers.
5. To the first beaker of analyte, add eight drops of phenolphthalein indicator, to the second add eight drops of bromothymol blue indicator. If you have time to do a third titration, use the methyl orange indicator so you can see a complete comparison. As you are doing your titration, ***be sure to note the pH at which the indicator changed color.***
6. Connect your pH probe to the pH input on the back of the **MicroLAB** interface and press the **Power On** button.
7. Open the **MicroLAB Experiment program** in the normal manner, select the experiment named **pH.Time.exp** from the **Titration** tab, then press **Enter** to open the program. Click on **pH** in the **Variables View** and recalibrate your probe with the pH 4, 7 and 10 buffers supplied. Be sure to rinse the pH probe with distilled water after each buffer and before you place it in your analyte solution. Between titrations, the probe should be stored in the pH 7 buffer, then rinsed well with distilled water before inserting into your titration beaker.
8. Place a stirring bar in the 250 ml beaker and position the beaker on top of the stirrer. Place the pH probe in the analyte solution at the side of the beaker and the buret above it to drop the titrant directly into the analyte. Now remove this beaker assembly and temporarily replace it with a "waste solution" beaker to get the titration started.
9. Set the drop rate at two to four seconds per drop. When the drop rate is properly set, while one student quickly replaces the "waste beaker" with the "titration beaker," the other student clicks the start button so that the program has started at the instant the first drop falls into the "titration beaker." Immediately turn on the magnetic stirrer and set it to a moderate rate.
10. Be sure to give a descriptive name for each titration when asked for it in the program. ***Do not use data,*** as that will be overwritten with new data each time.
11. The pH vs. Time data will immediately appear in **Graph, Digital Display** and **Spreadsheet Views**.
12. While one partner continues to maintain the titrant level between 0 and 10 with the wash bottle of 0.100 M NaOH, ***the other partner should watch for the change of color of the indicator. Be sure to record on the report sheet the pH and the time at which the indicator changes color at the time that the color change occurs.***
13. Continue the titration until the curve has flattened out at the top. Then turn off the stopcock (turn it

perpendicular to the length of the buret) and stop the data collection by clicking on "**Stop.**"

14. Repeat the entire process from step 10, using bromothymol blue in the second titration and methyl orange as the indicator for the third titration, if there is time.

***Be sure to use a different file name for each of the titrations and record this on your report sheet so you can recover them for printing and calculating later.***

Discard all titration solutions into a large beaker. At the end of the experiment, this should be neutralized and flushed down the drain with large amounts of water. Be sure solids are **NOT** washed into the sink. Any solids from this experiment may be disposed of in the container so labeled.

### DATA ANALYSIS

1. When the titrations are completed, reload each of your data files in turn into the *MicroLAB* program and perform the following analyses.
2. Click on the **Analysis** button, click on **Plot a Derivative**, use the current variables, **pH** and **Time**, click **OK**, The **Derivative** of pH vs. Time will appear on the **Graph View** in a fine line in red.
3. In the **Variables View**, scroll down to **Analysis (pH vs. Time)**, and "click drag" the **f'(x)** to the **Y2** Axis. This will show the derivative in blue, and with data points.
4. Now **Right Click** on the **f'(x)** line in the **Variables View** and click on **Hide this Derivative** to remove the red derivative line.
5. **Print** your graph. **DO NOT** print your data table, as that will require too much time and paper. However, using the scroll button on the right of the **Spreadsheet View**, scroll up and down to examine how the data in the table is changing.
6. Be sure to add your names or initials and a descriptive title to each graph that you print and submit. (The title should be added to the graph *before* you print.)

# Introduction to Acids, Bases and Titration

## INSTRUCTOR RESOURCES

### The CCLI Initiative

#### Learning Objectives

- introduce the nature of acids and bases.
- introduce acid-base indicators, e.g., litmus, wide range indicator papers and specific titration indicators.
- Introduce various acid/base reactions
- introduce titration as a means of determining the amount of an acid or base present.

#### Procedure Overview

- Students will explore the nature of acids and bases by observing the effect of buffers in the range of pH 2 to 11, on various indicator papers and solutions.
- Acid reactions are then explored using the common acids HCl, H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub> with substances such as marble chips, zinc, paper clips, aluminum, egg albumin and copper.
- The reactions of the common bases NaOH, Ca(OH)<sub>2</sub> and NH<sub>4</sub>OH are then explored with these same substances.
- Finally, a simple titration is performed using NaOH as the titrant and acetic acid as the analyte and utilizing the *MicroLAB* Interface to get a better understanding of a titration curve.

**INTRODUCTION TO ACIDS, BASES AND TITRATION**

**Report Sheet**

**Acidity and indicators**

1. Indicate how well the Alkacid paper color matched the color chart on the vial for each of the following buffer solutions. Indicate chart color and whether the observed color is lighter, darker or the same as the chart.

**2      3      4      5      6      7      8      9      10      11      12**

**Alkacid:**

**Chart:**

**Observed:**

2. How accurate and precise is this method of determining the acidity of a solution?

3. Specific pH indicators

<b>Indicator</b>	<b>Acid pH Range</b>	<b>Acid Color</b>	<b>Base pH Range</b>	<b>Base Color</b>	<b>Transition Range</b>	<b>Color</b>
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**Phenolphthalein**

**Methyl Orange**

**Bromothymol Blue**

4. How did the transition pH and the transition ranges vary for the three indicators?

5. How would these indicators be helpful in doing a titration without a pH meter?

Name \_\_\_\_\_ Section \_\_\_\_\_ Date \_\_\_\_\_

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Report Sheet

#### B. ACID REACTIONS

##### MARBLE REACTIONS:

ACID/BASE

OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

##### ZINC REACTIONS

ACID/BASE

OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

Name \_\_\_\_\_ Section \_\_\_\_\_ Date \_\_\_\_\_

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Report Sheet

#### IRON WIRE REACTIONS

ACID/BASE

OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

#### ALUMINUM REACTIONS

ACID/BASE

OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

Name \_\_\_\_\_ Section \_\_\_\_\_ Date \_\_\_\_\_

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Report Sheet

#### EGG ALBUMIN (WHITE) REACTIONS (A protein)

ACID/BASE

OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

#### COPPER REACTIONS

ACID/BASE

OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Report Sheet (page 2)

#### Titrations

File name:           1st \_\_\_\_\_           2nd \_\_\_\_\_           3rd \_\_\_\_\_

Calibration:        1st \_\_\_\_\_           2nd \_\_\_\_\_           3rd \_\_\_\_\_  
(sec/5 mL)

Indicator used:     1st \_\_\_\_\_           2nd \_\_\_\_\_           3rd \_\_\_\_\_

pH at color change: 1st \_\_\_\_\_           2nd \_\_\_\_\_           3rd \_\_\_\_\_

1. If you did not have a pH meter with which to do the HOAc titration, which would be the easiest indicator to use? Why?
2. Examine the time vs. derivative data in the table to find the maximum of the derivative, which represents the equivalence point for that titration. From this maximum, determine the time of the equivalence point for each titration.

Time of equivalence point:   1st \_\_\_\_\_   2nd \_\_\_\_\_   3rd \_\_\_\_\_

3. Which of the three indicators would be the best to use for the HOAc titration? Why?
4. What happens to the pH of the analyte as the titration nears the equivalence point? Explain why this occurs?
5. Why does the titration curve flatten out at the top?

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Tips and Traps

1. This is a fairly long experiment. If the storeroom people have all the materials packaged in cartons per pair of students, and if the students have organized their work schedules, they will have no trouble finishing in the allotted time.
2. During the calibration of the pH probes, students may compare millivolt readings and become concerned that they may differ widely from system to system. Assure them that this is natural and is a function of differences in the nature of the pH electrodes. The values should be around  $0 \pm 100$  millivolts if the pH probe is good.
3. You may find a wide variation in the quality of electrodes depending upon the manufacturer and/or age of the electrode. Some electrodes will reach equilibrium quite rapidly, while others may take a while. All electrodes should be checked out in advance to determine if they obtain equilibration within about 5 to 10 seconds. If not, the electrodes should be treated with an enzymatic cleaner, then soaked in an electrode conditioning solution, both of which can be obtained from Markson.
4. It is not necessary in this experiment to know absolute pH, only relative pH, since we are only looking at changes in pH to determine the general shape of the titration curve. Therefore, as in item 2 above, allow a reasonable time for equilibration during the calibration procedure, and then use that same time interval for each measurement during the titration.
5. There is a tendency for students to leave the pH probe hanging out in the air between titrations. If the time between titrations is small, this probably won't matter much, but it is better to teach them good practice, and that is to store the probe in their buffer in a beaker placed at the back of the bench between titrations. There is less chance of breakage this way, and the probe is equilibrating at or near the equivalence point pH.
6. In between uses, all electrodes should be kept in a sealed container with the bulbs covered with the conditioning solution mentioned above.

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Sample Data (1)

#### Acidity and indicators

- Indicate how well the Alkacid paper color matched the color chart on the vial for each of the following buffer solutions. Indicate chart color and whether the observed color is lighter, darker or the same as the chart.

	2	3	4	5	6	7	8	9	10	11	12
<b>Alkacid:</b>	purple	red	dark orange	light orange	yellow	light green	green	dark green	green/blue	light blue	blue
<b>Chart:</b>	purple		orange		gold		green		green/blue		blue
<b>Observed:</b>	same	lighter	same	darker	same	lighter	same	darker	same	lighter	same

- How accurate and precise is this method of determining the acidity of a solution?

*This method can approximate pH to about one or two pH units, depending upon how good the individual is at extrapolating color.*

- Specific pH indicators

Indicator	Acid pH Range	Acid Color	Base pH Range	Base Color	Transition Range	Transition Color
<b>Phenolphthalein</b>	2 - 8	colorless	10 - 11	pink	8 - 10	It pink
<b>Methyl Orange</b>	2 - 3	red	5 - 12	yellow	3 - 5	orange
<b>Bromothymol Blue</b>	2 - 6	yellow	8 - 12	blue	6 - 8	green

- How did the transition pH and the transition ranges vary for the three indicators?

*Each indicator has a different transition pH, but the range over which each indicator made the color transition was 1.5 to 2 pH units.*

- How would these indicators be helpful in doing a titration without a pH meter?

*Choosing the proper indicator would allow the chemist to stop the titration at or near the equivalence point, i.e., where equivalent amounts of base and acid have reacted.*

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Sample Answers (2)

#### B. ACID REACTIONS

##### MARBLE REACTIONS:

##### ACID/BASE

##### OBSERVATIONS

6 M HCl            The instructor should evaluate these responses in terms of the known reactions and the students abilities

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

##### ZINC REACTIONS

##### ACID/BASE

##### OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

Name \_\_\_\_\_ Section \_\_\_\_\_ Date \_\_\_\_\_

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Sample Answers (3)

#### IRON WIRE REACTIONS

ACID/BASE

OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

#### ALUMINUM REACTIONS

ACID/BASE

OBSERVATIONS

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

Name \_\_\_\_\_ Section \_\_\_\_\_ Date \_\_\_\_\_

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Sample Answers (4)

EGG ALBUMIN (WHITE) REACTIONS (A protein)

ACID/BASE	OBSERVATIONS
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6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

COPPER REACTIONS

ACID/BASE	OBSERVATIONS
-----------	--------------

6 M HCl

6 M H<sub>2</sub>SO<sub>4</sub>

0.6 M H<sub>2</sub>SO<sub>4</sub>

6 M NaOH

6 M Ca(OH)<sub>2</sub>

0.6 M NaOH

6 M HNO<sub>3</sub>  
(Hood)

6 M NH<sub>4</sub>OH  
(Hood)

## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Sample Data (page 5)

#### Titrations

File name:	1st	<i>Phenol Titr</i>	2nd	<i>MethOr Titr</i>	3rd	<i>Brom Titr</i>
Calibration: (sec/5 mL)	1st	<i>53285</i>	2nd	<i>42848</i>	3rd	<i>43356</i>
Indicator used:	1st	<i>Phenophth.</i>	2nd	<i>Methyl Or.</i>	3rd	<i>Bromo.Blue</i>
pH at color change:	1st	<i>9.2</i>	2nd	<i>3.5</i>	3rd	<i>7.2</i>

1. If you did not have a pH meter with which to do the HOAc titration, which would be the easiest indicator to use? Why?

*The easiest indicator to use would be the phenolphthalein, because it changes very sharply from colorless to pink, giving a very clear end point.*

2. Examine the time vs. derivative data in the table to find the maximum of the derivative, which represents the equivalence point for that titration. From this maximum, determine the time of the equivalence point for each titration.

Time of equivalence point: 1st \_\_\_\_\_ 2nd \_\_\_\_\_ 3rd \_\_\_\_\_

3. Which of the three indicators would be the best to use for the HOAc titration? Why?

*The phenolphthalein would also be the best indicator to use in this titration because the transition pH is very near to the pH at the equivalence point of the titration.*

4. What happens to the pH of the analyte as the titration nears the equivalence point? Explain why this occurs?

*The pH of the analyte increases very rapidly as the titration nears the equivalence point because there is very little hydrogen ion remaining in solution, and therefore very small amounts of added base removes a relatively large percentage of the remaining hydrogen ion.*

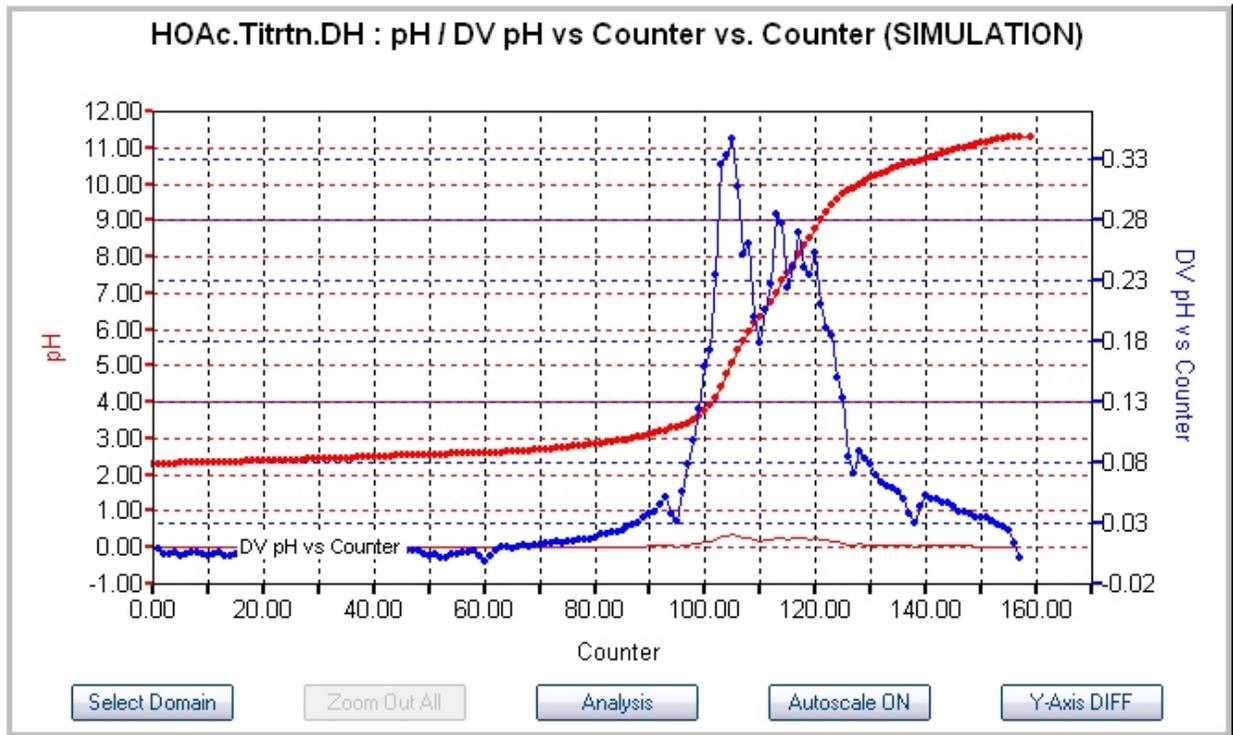
5. Why does the titration curve flatten out at the top?

*The titration curve flattens out at the top because base is being added to the solution without any acid to neutralize it. Thus, as larger amounts of base accumulate, the smaller the percentage change occurs as more base is added.*

# INTRODUCTION TO ACIDS, BASES AND TITRATION

## Sample Data (page 6)

This titration curve was collected using the photoelectronic counter, but it would look exactly like this just using Time as the X axis.



## INTRODUCTION TO ACIDS, BASES AND TITRATION

### Laboratory Preparation (per student station)

#### Equipment

1 - 50 mL buret with cork stopper	1 - buret clamp
3 - 250 mL beakers	1 - 25 mL beaker
1 - magnetic stirrer with stirring bar	1 - 50 mL beaker
50 - toothpicks	1 - ring stand
1 - pH probe with BNC connector	1 - twelve well, plastic well plate
1 - <i>MicroLAB</i> interface	1 tube of Alkacid for every pair of students

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#### Chemicals

The following chemicals will be required per student station

- 5 ml of buffer solutions 2 through 12
- 5 ml 6 M HCl
- 5 ml 6 M HNO<sub>3</sub>
- 5 ml 6 M H<sub>2</sub>SO<sub>4</sub>
- 5 ml 0.6 M H<sub>2</sub>SO<sub>4</sub>
- 5 ml 6 M NaOH
- 5 ml sat'd Ca(OH)<sub>2</sub>
- 5 ml 0.1 M NaOH
- 5 ml 6 M NH<sub>4</sub>OH
- 100 ml of 0.100 M NaOH solution, need not be standardized, but should be close.
- 30 ml of HOAc solution (HOAc is an abbreviation for acetic acid)
- 15 ml each of pH 4, 7 and 10 buffer solutions.
- 1 roll of Alkacid for every 2 pairs of students.
- Phenolphthalein, methyl orange and bromothymol blue indicators in dropper bottles.
- marble chips, crushed small
- granulated zinc, small pieces
- copper, small pieces
- small paper clip
- small pieces of aluminum
- 5 ml of egg white per group
- 100 mL of 0.100 M NaOH solution in a 250 mL wash bottle, need not be standardized, but should be close.
- 30 mL of ~0.100 M HCl solution
- 30 mL of ~0.100 M HOAc solution (HOAc is an abbreviation for acetic acid)
- 15 mL of each of pH 4, 7, and 10 buffer solutions per pair of students phenolphthalein, methyl orange and bromothymol blue indicators in dropper bottles per pair of students.

## Chemical Cautions

- **Hydrochloric Acid solution:** Toxic by ingestion and inhalation, strong irritant to eyes and skin.
- **Nitric Acid solution:** Very dangerous to eyes and skin. Severe tissue damage upon ingestion.
- **Sulfuric Acid solution:** Severely corrosive to eyes, skin and other tissue. Toxic, strong skin irritant. Powerful dehydrator causing blistering of the skin.
- **Acetic Acid solution:** Corrosive liquid, skin burns are possible, very dangerous to eyes.
- **Sodium Hydroxide solution:** Corrosive liquid, skin burns are possible, very dangerous to eyes.

The other chemicals are innocuous; however you should keep all chemicals away from eyes and mouth, wash hands after use and before leaving the laboratory, and use prudent laboratory practices at all times.

### Waste Disposal

- The acids and bases should be neutralized and then flushed down the drain with lots of water.

### DISPOSAL COLLECTION:

- A large container beaker in each hood labeled "WASTE HNO<sub>3</sub> SOLN."
- A large container on the reagent table labeled "WASTE SOLIDS" with about 100 ml of tap water in the bottom.

### WASTE DISPOSAL:

- The solids may be disposed of in the garbage after they have been rinsed several times with tap water to remove the acids.
- The HNO<sub>3</sub> in the fume hood beakers should be neutralized and then flushed down the drain with lots of water.