Experiment 7

Titration Curves of Strong and Weak Acids and Bases

Goals

To calibrate a pH probe. To become familiar with acid-base titration curves. To determine the concentration of an unknown acid.

Equipment and Materials

USB flash drive, 0.10 M ammonium hydroxide solution, 0.10 M sodium hydroxide solution, 0.0074 M hydrochloric acid solution, 0.0074 M acetic acid solution, a solution of hydrochloric acid (or acetic acid) with an unknown concentration, phenolphthalein solution, magnetic stir plate with magnetic stirrer, 50 mL burettes (2), 100 mL graduated cylinder, 25.00-mL volumetric pipette with pipette bulb.

Discussion

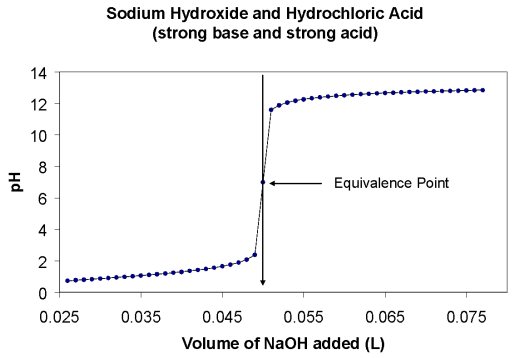
When a base is added to an acid a neutralization reaction takes place. The pH of the resulting solution depends on the concentrations of the acid and the base as well as the nature of the acid and the base. Strong acids and bases have a greater effect on pH than weak acids and weak bases at the same concentration.

A titration curve is a graph of how the pH changes as a titrant is

added to an analyte. The curve from the titration

of a strong acid (HCl) with a strong base (NaOH) is shown

in the figure. The reaction equation is:



$HCl\ (aq) + NaOH\ (aq) \rightarrow H_2O\ (l) + NaCl\ (aq)$

Sodium hydroxide is slowly added to the hydrochloric acid. At the beginning of the titration the pH is low because the solution contains HCl (a strong acid). As NaOH is added, the pH rises. The change in pH is slow at first. But near the equivalence point the pH changes rapidly. The equivalence point is the point on the curve where the number of moles of NaOH added is equal to the number of moles of HCl in the container. Past the equivalence point, the pH changes more slowly as NaOH is added.

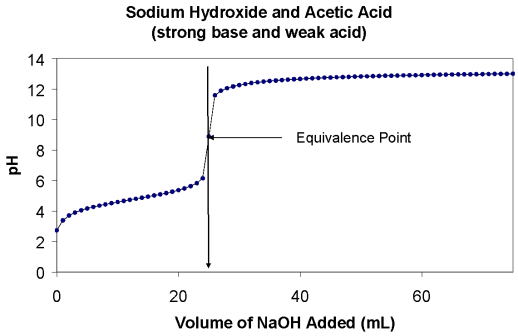
The Titration of a Weak Acid with a Strong Base

If we add NaOH to acetic acid (a weak acid) the resulting titration curve looks similar to the one for NaOH added to HCl. But there are some important differences.

The reaction equation is:

$$\underset{\text{acetic acid}}{HC_2H_3O_2\,(aq)} + \underset{\text{soidum hydroxide}}{NaOH\,(aq)} \rightarrow H_2O\,(l) + NaCl\,(aq)$$

Acetic acid is a weak acid (K_a = 1.8 x 10^{-5})



Notice that at the beginning of the experiment the pH is acidic. But less so than in the previous titration of HCl with NaOH. This difference is due to the difference in acid strength. The acetic acid is a weaker acid than HCl, so the initial

pH is higher. Another difference is the pH at the equivalence point. For the acetic acid reaction the pH at the equivalence point is greater than 7. In the reaction of HCl with NaOH the pH at the equivalence point was exactly 7.

In this experiment we will generate four titration curves (Part 2) for the following reactions:

$$\underset{\text{strong acid}}{HCl(aq)} + \underset{\text{strong base}}{NaOH(aq)} \rightarrow H_2O(l) + NaCl(aq)$$

$$\underset{\text{weak acid}}{HC_2H_3O_2(aq)} + \underset{\text{strong base}}{NaOH(aq)} \rightarrow NaC_2H_3O_2(aq) + H_2O(l)$$

$$\underset{\text{strong acid}}{HCl(aq)} + \underset{\text{weak base}}{NH_3(aq)} \rightarrow NH_4^+(aq) + Cl^-(aq)$$

$$\underset{\text{weak acid}}{HC_2H_3O_2(aq)} + \underset{\text{weak base}}{NH_3(aq)} \rightarrow NH_4^+(aq) + C_2H_3O_2^-(aq)$$

We will then (Part 3) use the equivalence point of the titration curve for the reaction of HCl with NaOH to determine the molarity of an HCl solution.

Procedure

SAFETY PRECAUTIONS

Safety glasses are required for this experiment. Hydrochloric acid can burn skin and should be handled with care. Sodium hydroxide can damage skin and should be handled with care. The glass electrode is a fragile device and should be handled with care.

Students work in pairs throughout this experiment.

Part 1: pH electrode calibration. We begin by calibrating the pH probe. The probe is potentiometric, meaning it measures electrical potential (voltage) to determine the pH of a solution. The device measures voltage, which is dependent on the pH of the solution the probe is placed in. We use pH standards (available from chemical supply companies like Fisher Scientific) to have known pH values and then the device measures the voltage for each pH solution. Common pH standards are 4.0, 7.0, and 10.0. Once the device has two values of pH and voltage, it assumes a linear relation between voltage and pH. We sometimes refer to this as a straight-line, or linear, calibration. We will use a pH probe and software for this procedure.

1. Double click on the **Vernier Graphical Analysis** to open it.
2. Click on the pH displayed reading located in the right lower corner and select **Calibrate**.
3. Choose to perform a Two-point calibration
4. Rinse the pH electrode with distilled water and blot dry it with Kimwipes.
5. Place pH electrode in the first buffer solution and **Enter the first known value**. Before pressing **Keep** make sure that the **Relative stability reading** displays a steady value. Press **KEEP**.
6. Rinse pH electrode with distilled water again before it is placed in the second buffer solution. Blot dry the electrode with Kimwipes.
7. Place the electrode in the second buffer solution and **Enter the second known value**. Make sure that the **Relative stability reading** displays a steady value. Press **KEEP**.
8. Select **APPLY.**
9. Leave pH electrode in the last calibration buffer solution to obtain a live reading of the calibrated electrode. The pH electrode reading should be within ± 0.02 pH units of the nominal pH value of the buffer solution.

Part 2: Shapes of Acid-Base Titration Curves

You will do four titrations in this part, with all curves plotted on the same chart. By plotting the titration curves together we can note the similarities and differences between weak and strong acids and bases. The shapes of these four curves will be compared and contrasted. In each case the basic solution will be added to the acidic solution. The four titrations are:

1.	sodium hydroxide + hydrochloric acid	NaOH + HCl
2.	sodium hydroxide + acetic acid	NaOH + HAc
3.	ammonium hydroxide + hydrochloric acid	NH_3 + HCl
4.	ammonium hydroxide + acetic acid	NH_3 + HAc

1. Clean two 50.00 mL burettes using tap water and then distilled or deionized water.
2. Rinse one of the burettes with a few milliliters of 0.10 M sodium hydroxide solution. Fill the burette above the 0.00 mL mark with the sodium hydroxide solution. Allow some of the solution to pour through the burette so that you can remove air bubbles at the bottom. Once the air bubbles have been eliminated you can let more solution out until the liquid is at the 0 mL level. You will use this buret for the first and second titrations.
3. Rinse the second burette with a few milliliters of 0.10 M ammonium hydroxide solution. Fill the burette to the 0.00 mL mark with the ammonium hydroxide solution. You will use this buret for the third and fourth titrations.
4. Using a graduated cylinder add 100 mL of 0.0074 M hydrochloric acid solution to a clean and dry 250 mL beaker. Add two drops of phenolphthalein indicator to the beaker. Place a magnetic stir bar to the solution and place the beaker onto a magnetic stirrer. Turn on the stirrer so that the solution is stirred without splashing onto the sides of the beaker.
5. Place the pH probe into the solution and clamp the probe to the ring stand. The pH should be between 2 and 3.
6. Setup the Graphical Analysis software for the experiment. To enter the experiment title, change pH range and extend the data collection time click on the **Graph Tools** located in the left lower corner and select **Edit Graph Options**. Add a graph title in the designated space. In the **y-axis range,** enter the pH range from 0 to 12 and in the **x-axis range** change data collection time to 500 seconds.

7. To enter the data collecting parameters, select Mode. The **Data Collecting Setting** will appear. Change the sampling rate to 0.5 samples/second and enter 500 seconds in the **End Collection** field. Press **DONE**.
8. Start the first titration: sodium hydroxide + hydrochloric acid. Begin collecting data by pressing **COLLECT**.
9. Slowly open the stop cock valve on the burette until the sodium hydroxide solution is flowing at about one drop per second. It may take a few seconds to get the drop rate right.
10. Watch the titration and record how long it takes for the solution to turn pink. Be sure to record that time and the pH at that time in the data sheet. You will need it for your report.
11. Allow the titration to continue until the experiment stops at 500 seconds.
12. To label the curve, select **Graph Options**, choose **Add Annotation.**
13. BE SURE TO SAVE THE FILE AFTER EACH TITRATION AS A .AMBL FILE.
14. To display all collected curves, click on the **pH** located on the y-axis. In the **Plot Manager**, click each **Data Set** to activate the graphs.
15. Refill the buret to the zero-milliliter mark before doing the next titration.
16. Continue with the second titration, acetic acid and sodium hydroxide. Here you will also use 100 mL of acetic acid solution. Be sure to save the file after each titration.
17. Continue with the final two titrations, hydrochloric acid with ammonium hydroxide, and acetic acid with ammonium hydroxide. You will need to refill the buret with more ammonium hydroxide solution in between these two titrations. Each time use 100 milliliters of the acid solution.
18. By the end you should have four titrations, each labeled in a chart, and time and pH of color change for each titration on the data sheet.
19. You can export the file as a pdf and print out the chart from a pdf reader.

Part 3: Potentiometric Acid-Base Titration

Here you will titrate sodium hydroxide solution against an acid solution, either hydrochloric acid (a strong acid) or acetic acid (a weak acid). Check with your instructor to know which acid you will use.

1. Use a Pipet Pump and the 25-mL volumetric pipet to add 25.00 mL of the acid solution into a 250-mL beaker. Enter this volume in the Data and Results Table.
2. Add 100 mL of distilled water to the beaker. Place a Teflon-coated magnet in the beaker.
3. Use a utility clamp to suspend a pH sensor on a ring stand. Position the pH Sensor in the acid solution and adjust its position at one side of the beaker so that its end is about ¼" above the teflon-coated magnetic stir bar.
4. Rinse the 50-mL buret with a 3–5 milliliters of the about 0.1 M NaOH solution. Use a buret clamp to attach the buret to the ring stand. Fill the buret a little above the 0.00-mL level of the buret with about 0.1 M NaOH solution. Drain a small amount of NaOH solution into a waste beaker so it fills the buret tip and leaves the NaOH at the 0.00-mL level of the buret. Examine the tip of the buret carefully to be sure that there are no air bubbles in it. If there are, let the NaOH solution drip into the waste beaker as fast as possible until you have swept out the bubbles. Refill the buret and re-adjust the volume to the 0.00-mL mark. Record the precise concentration of the NaOH solution in your data table. Dispose of the waste solution into the sink.
5. Prepare Graphical Analysis for data collection. Click on **Mode** to open the **Data Collection Settings**.
6. In the **Mode** window, select "Event Based".
7. Make sure that "Events with Entry" is chosen.
8. In the "Event Name" enter the name of the base.
9. In the "Units" window, enter mL as a unit of volume and press **DONE**.
10. Rescale the pH axis from 1 to 13 and the volume axis from 0 to 40 mL, click on **Graph Tools** icon and select graph options.
11. In the **Graph Options** window select **Manual Scaling** for both pH and Volume axes.
12. To collect the data with the appropriate precision, click on the triple dots displayed in the **Data Set 1** and select the **Column Options**. In the **Displayed Precision** select 2 for **Decimal Places**. Press **APPLY**.

13. Before adding NaOH titrant, start the magnetic stirrer, click **Collect** and monitor pH for 5-10 seconds. Once the displayed pH is stabilized, click **Keep**. In the entry box, type "0.00" as the NaOH volume. Press **KEEP POINT**.
14. Add the first increment of NaOH titrant to raise pH about 0.2 units above the original value. When the pH reading is stabilized, click **Keep** again. Enter the buret volume reading to the nearest 0.01 mL. Press **KEEP POINT**.
15. Continue adding NaOH solution in increments that raise the pH by 0.2 units and enter the buret volume reading after each increment.
16. When a pH value of approximately 3.5 is reached, change to a one-drop increment. Enter a new buret volume reading after each drop is added. To get the best-looking titration curve, each increment should be the same size-namely one drop.
17. After a pH value of approximately 10.5 is reached, again add larger increments of NaOH that raise the pH by about 0.2 pH unit, and enter the buret volume reading after each increment.
18. Continue adding NaOH solution until you have added about 40 mL of titrant in total.
19. When you have finished collecting data, click **STOP**. Dispose of the beaker content into the appropriate disposal container.
20. To display the first derivative, click on the triple dots above the pH column, select "Add Calculated Column", click on the **INSERT EXPRESSION,** choose "1st Derivative (X,Y)" and click **APPLY**.
21. To display the 1st Derivative graph, click on the pH located on the y-axis of the graph and activate the "First Derivative" button. If the graph is not completely displayed, click on the zoom icon located on the bottom left corner of the Vernier Graphical Analysis.

Name_____ Date_____ section_____L

Data Sheet

Experiment 7: Titration Curves of Strong and Weak Acids and Bases

Part 1. Calibration of pH glass electrode

pH of first buffer solution _____

pH of second buffer solution _____

Part 2. Shapes of acid-base titration curves

Titration	**initial pH**	**final pH**	**time of color change**	**pH at color change**
HCl + NaOH	_____	_____	_____	_____
HAc + NaOH	_____	_____	_____	_____
HCl + NH_4OH	_____	_____	_____	_____
HAc + NH_4OH	_____	_____	_____	_____

Part 3. Quantitative Determination of the Concentration of HCl or Acetic Acid

Name_____ Date_____ section_____L

Volume of acid solution added to beaker	mL
Concentration of NaOH (obtained from your instructor)	M
Volume of NaOH solution added at the equivalence point	mL
Moles of NaOH added at the equivalence point moles of NaOH = (molarity)(volume)	mol
Moles of acid in the beaker moles acid = moles of base at equivalence point	mol
Calculated concentration of acid (reported to 4 significant figures)	M
*For acetic acid pK_a from curve	
For the instructor's use only	% error =

*The pK_a of a weak acid is equal to the pH at one-half the equivalence point. Find the volume of the equivalence point, divide that volume by 2. Then find the pH at that volume from the data table on the computer screen.

Show your calculations

Table 1: Acid Base Indicators

Indicator	pH range	color change	Preparation
methyl violet	0.0 - 1.6	yellow to blue	< 0.1% in water
crystal violet	0.0 - 1.8	yellow to blue	< 0.1% in water
malachite green	0.2 - 1.8	yellow to blue-green	0.1% in water
erythrosine, disodium	2.2 -3.6	orange to red	0.1% in water
bromophenol blue	3.0 - 4.6	yellow to blue	basic solutions
congo red	3.0 -5.0	blue to red	0.1% in water
methyl orange	3.2 -4.4	red to yellow	
ethyl orange	3.4 - 4.8	red to yellow	
ethyl red	4.0 - 5.8	colorless to red	
bromocresol green	4.0 - 5.4	yellow to blue	
alizarin red S	4.6 - 6.0	yellow to red	
methyl red	4.8 - 6.0	red to yellow	
p-nitrophenol	5.4 - 6.6	colorless to yellow	
alizarin	5.6 - 7.2	yellow to red	
	11.0 - 12.4	yellow to red	
brilliant yellow	6.6 - 7.8	yellow to orange	
phenol red	6.6 - 8.0	yellow to red	
cresol red	7.0 - 8.8	red to yellow	
thymol blue	8.0 - 9.6	red to yellow	
phenolphthalein	8.2 - 10.0	colorless to pink	
thymolphthalein	9.4 - 10.6	colorless to blue	
alizarin yellow R	10.1 - 12.0	yellow to red	
2,4,6-trinitrotoluene	11.5 - 13.0	colorless to orange	

Laboratory Report

1. For Part 2, print the graph with all four titration curves on one chart. Be sure each curve is properly labeled with text annotation to indicate what it is showing. Be sure the axes are properly labeled.
2. Fill out both data pages with the data and calculations taken during the experiment.

3. Answer the following questions for Part 2, being sure to show your work step by step. Answers will not receive credit unless the steps are shown explicitly:

    - Calculate the expected pH of the hydrochloric acid solution used in the experiment.
    - Calculate the expected pH of the sodium hydroxide solution used in this experiment.
    - Calculate the expected pH of the acetic acid solution used in the experiment. You'll need to look up the K_a value for acetic acid.
    - Calculate the expected pH of the ammonium hydroxide (ammonia) solution used in the experiment. You'll need to look up the K_b value for ammonia.
    - Refer to Table 1 on page 10, a list of acid-base indicators. Use the list to find at least two indicators that might be effective for the titration of sodium hydroxide with acetic acid. Explain the logic of your choices.

4. For Part 3, show the calculations for finding the concentration of the acid (HCl or acetic acid). If you worked with acetic acid, also show the calculations for the pK_a from the data set.