Don’t forget to bring a sample of a beverage you drink often!!

Pre-lab: Prepare the typical pre-lab sections (purpose, data tables). Data for this lab will include all acid solution concentrations and pH’s, initial and final volume readings for three titrations each for the “A” acids, and at least 20 pH values of beverage samples.

Introduction

In this experiment you will learn the difference between strong and weak acids, how to measure pH and calculate the hydronium ion concentration for an acidic solution, how to calculate percent ionization, how to titrate acidic solutions with a base, how to use titration data to calculate an unknown molarity, and how to predict a titration volume.

Strong Acids - An acid is called “strong” if almost all the acid molecules lose an H⁺ (proton) in water. The reaction of a strong acid in water proceeds forward to products almost completely, so no equilibrium is set up because the reverse reaction does not occur to any appreciable extent. For all practical purposes the reaction is over. We say the ionization of a strong acid is about 100% because almost all of the original acid, HA, ionizes in water.

\[
HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)
\]

Weak Acids - An acid is called “weak” if only a few of the acid molecules lose an H⁺ (proton) in water. The reaction of a weak acid in water proceeds forward to products until it reaches equilibrium, which is when the reverse reaction is going at the same speed as the forward reaction. An equilibrium reaction is never over because it continues to proceed forward and backward at the same rate forever if not disturbed. We say the ionization of a weak acid is considerably less than 100% (often less than 10%) because few of the original acid molecules, HA, ionize in water; thus, few products are formed.

\[
HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)
\]

Percent Ionization - A useful measure of the strength of an acid is its percent ionization, which is given by:

\[
\text{% Ionization} = \frac{[H_3O^+]}{[HA]_{initial}} \times 100\%
\]

The hydronium ion concentration is measured after the acid has been mixed with the water, so it represents the final or equilibrium concentration of H₃O⁺(aq). The initial acid concentration is the total acid concentration before the reaction takes place. Note: Acids react almost immediately with water so all your measurements will represent values after reaction.

Relationship Between pH and Hydronium Ion Concentration

The pH of a solution is defined as: pH = -log [H₃O⁺]. Because the pH is the negative log of [H₃O⁺], the pH decreases as the [H₃O⁺] or acidity increases. Rearranging this equation to solve for [H₃O⁺], we obtain [H₃O⁺] = antilog (-pH) = 10⁻ᵖʰᴴ.

Titration of acids with a strong base - When any acid reacts with a strong base the neutralization reaction will proceed forward to completion. Even weak acids react completely with a strong base. Strong bases produce hydroxide ions in aqueous solutions and hydroxide ions have such a high affinity for protons that they will pull all the protons off weak acids or strong acids. Water is a much weaker base than hydroxide so water molecules are not able to pull all the protons off weak acids. The neutralization reaction for a acid-base titration can generally be written:

\[
HA(aq) + NaOH(aq) \rightarrow NaA(aq) + H_2O(l)
\]

Stoichiometry can be used to perform many titration calculations including calculating the unknown molarity for an acid or base solution as well as determining the volume of titrant needed to reach the equivalence point. (Review sections 6.10 and 6.11 in the McMurry/Fay textbook.)
**Techniques:** Before beginning a titration be sure to condition the glassware to be used. Refer to the handout on Conditioning and Using Glassware for Titrations.

**Titration** - When acid and base solutions are titrated together, one is typically put into an Erlenmeyer flask with its volume measured accurately, and the other (called the titrant) is put into a buret. A magnetic stir bar is used to stir the solutions together as the titrant is slowly added. The titration is over once the solution in the flask is neutralized, which is when exactly enough moles of the titrant have been added to react with all of the moles present in the flask. This is the **equivalence point.** Thus the resulting solution has no reactants left, nothing is in excess (there is no acid left and no extra base), and only products are present. If you add too much titrant, then it is still present in the beaker unreacted and the titration must be redone.

**Important Note:** In this lab the acid will be in the flask and the base will be the titrant. At the **equivalence point** just enough base has been added to react with all the acid, thus the resulting solution contains only products (there is no limiting reagent and nothing is in excess, the exact “stoichiometric” amounts have been added). However, at this point the solution is still colorless. To observe a visible change we will use phenolphthalein indicator. This indicator changes from colorless (in acidic solution) to pink (in basic solution) after a drop of excess base is added. Thus the color change actually occurs after the equivalence point has been passed. The **endpoint** is when the indicator changes color, and occurs when a little excess base is present. A good titration is when the endpoint is as close as possible to the equivalence point, thus the solution is barely pink, and so only a tiny amount of excess unreacted base is present. The titrant volume at this endpoint is so close to the equivalence point that the measurement is good enough to approximate the equivalence point.

**Procedure**

I. **Measuring pH values of acids and beverage samples:**

1. Condition the 5.00 mL pipet and 25.00 mL buret following the steps in the “Conditioning and Using Glassware for Titrations” handout.

2. Obtain a magnetic stir bar, a pH probe, a buret, and a 5.00 mL pipet. Obtain and assemble the laptop computer and LabPro system as directed on the handout for using Computers with LabPro sensors. Attach the probe to the CH 1 outlet on the LabPro unit. **Allow some space between your titration equipment and the computer!**

3. Double-click on the “Shortcut to GCC Chm Labpro” folder on the desktop. Double-click to open the “Intro Acids” file. When Logger Pro launches, you should see a reading for the pH sensor displayed on the screen. Calibrate the pH probe following the steps in the “Using and Calibrating a pH Probe” handout.

4. Obtain approximately 30 mL of nitric acid A, 20 mL of nitric acid B, 30 mL of glycolic acid A and 20 mL of glycolic acid B in four different clean, dry and labeled 50 mL beakers. **Record the exact concentrations for all four acids in your lab notebook.**

5. Measure the pH of all four acid solutions and your beverage samples as described in the Techniques section. **Record the pH readings in your lab notebook.** (Keep the nitric A and glycolic A acids for part II.)

II. **Titrating the “A” acid samples (nitric A and glycolic A):**

6. Obtain approximately 100 mL of sodium hydroxide in a clean, dry 150mL beaker. Prepare the buret to be filled with this solution as described under the Techniques section. **Record the initial buret reading in your lab notebook.** (Save the rest of the solution to refill the buret later.)

7. Pipette 5.00 mL of the nitric acid A solution into a thoroughly rinsed 125 mL Erlenmeyer flask (as described in the handout) and add 2 drops of phenolphthalein. Add approximately 15 mL of DI water.

8. Add the stir bar to the flask by holding the flask at an angle so the bar slides into the solution and does not splash, and place on the stir plate. Turn the stir knob until the solution is stirring at a brisk pace without splashing. Position the buret so that the tip is just inside the flask. Titrate the solutions to a light pink endpoint. **Record the final buret volume in your lab notebook.**

9. **Repeat steps 6, 7, and 8 for two more trials (a total of 3 trials) with nitric acid A solution.** (Note: We are not titrating acid B.)

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10. Pour your nitric acid and titration solutions down the drain with plenty of water. Rinse your beakers and flasks thoroughly with tap water. Give them a final rinse with DI water. The flasks may be reused wet.

11. Repeat steps 6 – 10 but use glycolic acid A (a monoprotic acid with the formula \( \text{HC}_2\text{H}_3\text{O}_3 \)) instead of nitric acid. Refill the buret with sodium hydroxide to just below the zero mark.

12. Have your instructor initial the pH values you recorded for the four acid solutions and your titration data for the six trials before disassembling the equipment.

13. When your group is all done with the lab, make sure you have recorded all the beverage pH values from your class in your lab notebook. Instructors will keep a running list of all classes’ values on the side board.

**Waste Disposal:** Pour the solutions in the sink with running water. Cleanup: Rinse all glassware with lots of tap water then a final rinse with DI water. Wipe your entire bench down with a wet paper towel.

**Calculations:** Complete the following calculations (numbers 1 – 7 below) in your lab notebook! Be sure to number and clearly label each calculation. Include units for all calculations and watch significant figures!

1. (10 pts) Use pH values to calculate the hydronium ion concentrations for all four acid solutions. Then using the exact initial concentrations, calculate the percent ionization for each acid. Show example calculations for \([\text{H}_3\text{O}^+]\) and % ionization for just the nitric acid A solution. Include a summary table in your notebook for all 4 solutions. Your column labels should be solution, pH, \([\text{H}_3\text{O}^+]\), \([\text{HA}]_{\text{initial}}\), and % ionization.

2. (7 pts) Write the balanced equation for the reaction between aqueous nitric acid and aqueous sodium hydroxide. Using the data from the nitric acid titrations, calculate the concentration of the sodium hydroxide for each of the three trials. Show the calculation set-up (with units) for one of the trials, but make sure to report the calculated concentration of sodium hydroxide for all three trials. Do NOT use \(M_1V_1 = M_2V_2\) as that is for dilutions not titrations.

3. (7 pts) Write the balanced equation for the reaction between aqueous glycolic acid (\(\text{HC}_2\text{H}_3\text{O}_3\)) and aqueous sodium hydroxide. Using the data from the glycolic acid titrations, calculate the concentration of the sodium hydroxide for each of the three trials. Show the calculation set-up (with units) for one of the trials, but make sure to report the calculated concentration of sodium hydroxide for all three trials. Do NOT use \(M_1V_1 = M_2V_2\) as that is for dilutions not titrations.

4. (1 pt) Calculate the average concentration of sodium hydroxide from both sets of acid titrations (from questions 2 and 3). Note: you should only report ONE average concentration based on the calculated concentration of NaOH from all six trials.

5. (2 pts) Obtain the true value for the concentration of sodium hydroxide from your instructor. Write this value in your lab notebook (labeled!). Calculate the percent error using the equation below where the measured is your average NaOH concentration from #4 and actual is the true value:  

   \[ \% \text{ error} = \left( \frac{\text{measured} - \text{actual}}{\text{actual}} \right) \times 100 \]

6. (3 pts) Using the average sodium hydroxide concentration calculated in question 4, calculate the volume needed to neutralize 5.00 mL of the nitric acid B. Use the exact nitric acid B concentration in your set-up. Do NOT use \(M_1V_1 = M_2V_2\) as that is for dilutions not titrations.

7. (3 pts) Using the average sodium hydroxide concentration calculated in question 4, calculate the volume needed to neutralize 5.00 mL of the glycolic acid B. Use the exact glycolic acid B concentration in your set-up. Do NOT use \(M_1V_1 = M_2V_2\) as that is for dilutions not titrations.

**You may neatly hand-write your conclusion in your lab notebook.** Compare properties (pH, % dissociation, etc.) of strong versus weak acids. Report your average NaOH concentration and % error. Also discuss 3 sources of error and how they would have affected your results.
Discussion Questions (answer these on the handout pages): For these questions, use complete sentences with proper grammar and English. Your calculations and conclusion can be completed in your lab notebook.

1. (2 pts) Write the balanced reaction for hydrochloric acid in water.

2. (2 pts) Write the balanced reaction for benzoic acid (HC\textsubscript{7}H\textsubscript{5}O\textsubscript{2}) in water.

3. (5 pts) Circle the acid solutions below that would be considered “strong.”
   a. perchloric acid
   b. glycolic acid
   c. sulfuric acid
   d. ascorbic acid
   e. nitric acid

4. (3 pts) Explain your choices in number 3 above. What is the difference between a strong acid and weak acid?

5. (4 pts) Put the four acid solutions (label by acid name and A or B; i.e., nitric B) in order of decreasing pH.

<table>
<thead>
<tr>
<th>Highest pH</th>
<th>Next</th>
<th>Next</th>
<th>Lowest pH</th>
</tr>
</thead>
</table>

Explain how glycolic acid can have a lower pH and be more acidic than nitric acid.

(16 pts) For the following statements, circle the bold and italicized phrase that correctly completes each one.

6. A strong acid has a high / low percent dissociation.

7. A weak acid is a nonelectrolyte / weak electrolyte / strong electrolyte.

8. A strong acid is a nonelectrolyte / weak electrolyte / strong electrolyte.

9. At the same molarity, a strong acid will have a higher / lower / same pH than a weak acid.

10. As concentration of a weak acid decreases, the pH increases / decreases / is constant.
11. As concentration of a weak base decreases, the pH increases / decreases / is constant.
12. As concentration of a weak acid decreases, the percent dissociation increases / decreases.
13. As the strength of an acid increases, the conductivity increases / decreases / is constant.
14. (3 pts) Consider your answers to calculations number 6 and 7 on page 3 (not questions 6 and 7). Explain why the sodium hydroxide volumes needed to titrate the dilute nitric acid and glycolic acid are nearly the same or why they are very different. Should a stronger acid require more sodium hydroxide to neutralize it than a weaker acid? Explain why or why not.

15. (6 pts) Identify what is wrong with each of the following pictures (there might be 1, 2, or more errors):

16. (4 pts) On the line below, write the equation for the reaction between nitric acid and sodium hydroxide. In the left beaker, draw a picture of the equivalence point for the titration between nitric acid and sodium hydroxide. In the right beaker, draw the endpoint for the same titration. (Hint: determine how many molecules or ions of each reactant and product should be shown in each drawing if you start with 5 molecules of acid. See the note on page 2 about these regions.)

Equation: _____________________________

equivalence point

endpoint