Borrowed From Lowell High School AP Chemistry — Spring 2021
Titration of a Weak Acid — Acetic Acid in Vinegar

Introduction:

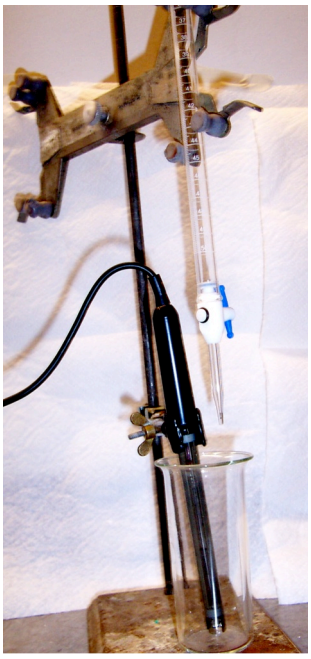
Vinegar is one of a standard household substance that you have no doubt used. We add vinegar to our salad dressing and other dishes. Sometimes we take advantage of its acidic properties and use it for a cleanser. Distilled white vinegar contains only acetic acid and water. You will determine the amount of acetic acid in vinegar by titration.

Purpose:

To determine the molarity, % by mass, pK_a, and K_a of acetic acid, CH_3COOH, in distilled white vinegar.
To be able to visually represent what is occurring during the acid-base titration on a molecular level.

Background Information:

Analysis by titration is a very useful and commonly used technique in chemistry. The sample to be analyzed is usually a solution in a tall beaker. A reagent is carefully added to the sample using a buret. A buret is like an upside down graduated cylinder. The reagent is added until the entire sample has reacted. Knowledge of the volume of reagent added and the concentration of the reagent allows us to calculate the amount of the substance in the sample that we are interested in. How will you know when to stop adding the reagent in the buret? An indicator and a pH probe (see figure at right) are used to determine the equivalence point or endpoint of the reaction.



The equivalence point is where the number of moles of H^+ in the sample is equal to the moles of OH^- you have added with the buret. From this you can get the moles of acetic acid in the sample of vinegar you are titrating. The equivalence point of the reaction is determined by the use of a visual indicator and a probe. For this titration the indicator is phenolphthalein, which changes from colorless at pH = 8 to red at pH = 10. Our sample, vinegar, is an acid, with pH less than 7. Our reagent will be sodium hydroxide, a base. The vinegar is placed in a tall 180 mL beaker with a few drops of phenolphthalein. The base is added in small portions from the buret so that the addition can be stopped when a permanent pinkish color appears.

You will determine the molarity, % by mass, pK_a, and K_a of acetic acid in vinegar from the following information:

- The volume of NaOH(*aq*) used to reach the equivalence point.
- The concentration of NaOH(*aq*) in moles/L (given).
- The stoichiometry of the reaction: $NaOH(aq) + HC_2H_3O_2(aq) \rightarrow H_2O(l) + NaC_2H_3O_2(aq)$
- Volume of vinegar and density of the vinegar.
- Molecular Weight of acetic acid.
- The pH of "pure" vinegar.
- The pH curve plotted as pH vs. mL of NaOH added.

Materials:

2 250-mL tall beakers
1 50-mL buret
Funnel
Distilled water
Computer
Cables
Graduated pipet
Distilled white vinegar
1 buret stand and holder
phenolphthalein
0.50 M NaOH(aq)
Interface
pH probe
LoggerPro 3.8.4 software

Procedure:

1. Cleaning and rinsing buret and beakers:
   Cleaning: Rinse the buret as follows. Obtain and clean a funnel and rinse the buret with tap water. Be sure to let the water run out of the tip. Rinse the buret twice with approximately 10.0 mL portions of 0.5 M NaOH(*aq*). Rotate the buret to make sure the sides are coated with the NaOH(*aq*). Let some of the NaOH(*aq*) out of the tip and pour the rest down the sink. Obtain two 180 mL tall beakers and use soapy water to wash the beakers. Finally, rinse the beakers with distilled water.

2. Filling buret with NaOH:
   Fill the buret to just above the 0.00 mL line with the NaOH solution. To eliminate bubbles from the tip, let a small portion of the NaOH(*aq*) solution out of the tip into a beaker until the solution is at the 0.00 mL line. Although starting at 0.00 mL is not necessary for the experiment, it makes calculating how much base you've added easier. Check that the tip of the buret is now filled and that there are no air bubbles in it. Record your initial volume of NaOH(*aq*) in the buret.

3. Filling beakers:
   Add about 10 mL of vinegar to each beaker using the graduated pipet so that you know the volume of vinegar added to the nearest 0.01 mL. Add about 25 mL of distilled water and four drops of phenolphthalein to each beaker as well.

4. Setting up pH probe:
   Carefully but firmly remove the pH probe from its buffer storage bottle without taking off the bottle's lid. Rinse the tip of the probe in tap water and remove drops of water by shaking a bit. Firmly clamp the probe in the beaker without tipping it all over and without getting in the way of the spinning stir bar.

5. Titration trial #1:
   Titrate the vinegar solution by adding NaOH(*aq*) from the buret into the beaker containing the vinegar. Use trial #1 to quickly find roughly where the endpoint is. For trial #1, do a relatively coarse titration with 1 to 2 mL increments of base added. For the pH sensor, click "Collect" when you are ready to start the experiment and click "Keep" each time the pH has stabilized and you want to record that value. In the edit box, type the **total** mL of base added so far and press ENTER. Stop the titration when a faint pink color appears and the probe indicates neutralization has occurred. This is the equivalence point. If the color is hot pink, redo the titration. Don't forget to "Keep" the pH value at 0.00 mL of base added!

6. Titration trial #2:
   a. Refill the buret.
   b. Repeat this titration for the second sample of vinegar. For trial #2, use your knowledge of where the endpoint is to go relatively quickly in the flat parts and much more slowly immediately before and after the endpoint.

7. pH of "pure" vinegar:
   Determine the pH of "pure" vinegar by rinsing the beaker and probe and then pouring enough vinegar into the beaker so as to cover the tip of the pH probe completely so it gets a good reading.

8. Clean-up:
   Rinse out everything. Put the pH probe back into its buffer storage bottle. To dry the buret, put it inverted in the buret clamp with the stopcock in the open position.

Calculations

Show all work for these calculations. Record final answers in an analysis table.

1. Write the net ionic equation for the reaction that occurred between the acetic acid and NaOH.

From your **titration**:

2. Determine the **moles of OH^-** from NaOH added to reach equivalence and the moles of acetic acid that were in your 10.0 mL vinegar sample.

3. Determine the concentration, in **mol/L, of acetic acid** in vinegar.
4. Determine the **% by mass of acetic acid** in vinegar. Assume the density of vinegar is 1.0 g/mL.

Using the molarity of acetic acid in vinegar from #3 above and the pH of pure vinegar from Procedure step #7, determine:

5. the $\mathbf{K_a}$ of acetic acid. Hint: what part of the ICE table does the pH tell you? Double hint: This is the same type of problem as Ch 14 pg 675 #67.
6. the $\mathbf{pK_a}$ of acetic acid from the K_a you just got.

Looking at your pH curve and taking the pH at the half-equivalence point (half-way to the equivalence point, when half of the acetic acid has been neutralized), do the following:

7. Write out the Henderson-Hasselbalch equation for acetic acid.
8. Write the Henderson-Hasselbalch equation for what it looks like at the half-equivalence point. Hint – think what must be true about the relationship between the concentrations of acetic acid and acetate anion.
9. From your pH curve at the half-equivalence point, determine the pK_a of acetic acid.
10. From this pK_a, calculate the K_a of acetic acid.
11. In the data table, list your values next to these accepted values and include the % difference.
    Acetic acid: $K_a = 1.8 \times 10^{-5}$, $pK_a = 4.74$
    Distilled white vinegar: 0.83 M acetic acid, 5% acetic acid by mass*

*Vinegar jug says 5% "acidity" which means 5 grams acetic acid per 100 mL of vinegar and to 2 sig figs, vinegar's density is 1.0 g/mL.

Data Analysis:

On your titration curve, draw 4 beakers which have inside them particulate representations showing what species are present in the solution at the following points in the titration:

1. no NaOH added
2. half-equivalence point
3. equivalence point
4. half-way past the equivalence point where excess NaOH has been added

Draw arrows to indicate which part of the titration curve each beaker represents. For beaker #1 (no NaOH yet added), assume you start with 8 molecules of weak acid HA. Hint: Choose from among these species for your particulate representations: HA, H_3O^+, A^-, OH^-, H_2O (only water molecules produced during reaction need to be shown).

Discussion Questions:

1. Why might it be important to rinse the buret so thoroughly in Step #1?
2. Describe how your concentration of acetic acid in vinegar would be off if you didn't realize you had added too much NaOH and had gone past the endpoint of the titration and recorded that volume of NaOH as the amount required to neutralize the acetic acid?
3. Why won't the equivalence point change when 25 mL of water was added to your 10.0 mL of vinegar in Step #3?
4. The density of vinegar is not exactly 1 g/L. But based on your % by mass of acetic acid in vinegar, why might it be a good approximation to use the density of vinegar as 1.0 g/mL in Calculations step #4?
5. What possible sources of error might have contributed to your numbers being off in the direction they are? For each proposed source of error, explain how it affected your data in that particular direction.
6. Using your data, explain the difference between the endpoint determined by your pH indicator and the equivalence point determined using the pH probe.

Conclusion:

Follow regular lab report guidelines for the conclusion.