
Objective

In this experiment we will investigate three titration curves of known acids and bases as well as that of an unknown. This analysis will yield two pieces of information about each acid. First, you will determine the molarity of each acid. Keep in mind that the molarities provided on the container of each known acid are approximate. Second, you will determine the K_a of each acid.

The molarity of each acid can be experimentally determined by titration with a base of known concentration. The number of moles of base added, the stoichiometric ratio of base to acid, and the *original* volume of acid are all needed in order to determine the molarity of the acid. This is something that you did in both Chemistry 120 and Chemistry 141.

The K_a of an acid can be determined by titrating the acid and determining the pH at a point half-way to the equivalence point. If the equivalence point occurs at a volume of 31.10 mL, then the so-called half equivalence point would be at a volume that is one half of this value (15.55 mL). Why is knowing the pH at this point helpful in determining the K_a of an acid? How are the concentrations of the protonated acid, HA, and conjugate base, A^- , related at this point in the titration?

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \quad \text{(EQ 4.1)}$$

Procedure

You will need to work in pairs to complete this experiment. You will use the Vernier computer interface to collect your data. It is important that each student be familiar with Vernier function as well as the data analysis tools available in the package. Each student will be required to perform the necessary calculations.

Each student pair is to titrate the acids listed below with 50mL of 0.1000M NaOH solution. *Make sure that you record the actual molarity of the NaOH.* It will be close to 0.1000M, but it most-likely won't be exactly 0.1000M.

1. 25.00mL of 0.100M HCl
2. 25.00mL of 0.100M HC₂H₃O₂
3. 25.00mL of 0.050M H₃PO₄

You will measure the pH of the solution as a function of the total mL of NaOH added.

After the first three titrations, *each* student will titrate 25.00mL of an unknown acid, obtained from the stockroom, with 0.1000M NaOH.

Using the Vernier Interface

You will be measuring the volume of NaOH added to the acid solution with a drop counter. Since we need actual volume in mL it is necessary to calibrate the drop counter.

Calibrating the Drop Counter

1. Place a clean, dry, *weighed* beaker underneath the drop counter and start measuring drops.
2. Open the stopcock on the separatory funnel to begin dropping the NaOH into the clean, dry, *weighed* beaker.
3. The computer will then begin recording the number of drops added to the beaker. After approximately 5-10mL of NaOH has been added to the beaker, close the stopcock on the funnel and stop the measuring of drops. Record the number of drops placed into the beaker.
4. Weigh your beaker again and determine the mass of NaOH in the beaker. Use a density value of 1.00g/mL for the NaOH solution to calculate the actual volume of NaOH per drop.

$$\frac{\text{Mass of NaOH}}{\# \text{ of Drops}} \times \frac{1 \text{ mL NaOH}}{1.00 \text{ g NaOH}} = \text{mL NaOH/drop NaOH} \quad (\text{EQ 4.2})$$

You will use this value for mL/drop to determine the total volume of NaOH added to your flask.

Collecting Titration Data

Once your drop counter is calibrated you are ready to calibrate your pH meter and collect your titration data.

1. Calibrate your pH electrode using two buffer solutions of known pH. Once the reading stabilizes, follow the directions on the screen.
2. You are now ready to collect pH as a function of number of drops added. Make sure that you prepare your solutions to be titrated, and begin measuring the pH as you add the NaOH dropwise. Collect data until you have reached 1.5 times the volume at the equivalence point.

Data Analysis

Plot four separate, full-page graphs. Each plot should include the following.

1. The equivalence point. You may locate this by performing the appropriate calculus (where the second derivative is 0). You may also locate it graphically. Make sure that you plot the volume of the NaOH on the x -axis, and not # of drops.
2. On the graph indicate the predominant species present in the titration beaker:
 - a. At the half equivalence point
 - b. At the equivalence point
 - c. At a point one and one half times beyond the equivalence point. You do not need to include water in this list.
3. Calculate the molarity of the titrated acid calculated from the location of the equivalence point and the concentration of the base used for the titration. Do this with as great a precision as your plot allows. Keep in mind that this is a Chemistry 120 type calculation. You have done one or two acid-base titrations prior to this; it is the same calculation as you have done in the past.
4. The pH of the system can be calculated from the appropriate equilibrium constant before, at, and past the equivalence point. For acetic and phosphoric acids, write out the equilibrium constant expression (along with its numerical value) that *could* be used to calculate the pH of the system before, at, and past the equivalence point. If, at any one of these points, the pH is determined primarily by the presence of a strong acid or base, then state this fact and do not try to give an equilibrium expression.
5. Report the experimentally determined K_a for the dissociation of the acid **using the pH at the half equivalence point of the titration**. Do this for the acetic acid, for K_{a1} and K_{a2} of the phosphoric acid *and for your unknown acid*.
6. In your "Data and Calculations" section of your lab report be sure to report your unknown number, the molarity of your unknown acid and its K_a value.

