

The determination of acid constant K_a



**ASSIGN
BUSTER**

The Determination of Acid Constant K_a Abstract When an acid or base dissolves in water, a certain percentage of the acid or base particles will break up, or dissociate, into oppositely charged ions. The Arrhenius theory defines an acid as a compound that can dissociate in water to yield hydrogen ions, H^+ , and a base as a compound that can dissociate in water to yield hydroxide ions, OH^- . The base sodium hydroxide, $NaOH$, dissociates in water to yield the required hydroxide ions, OH^- , and also sodium ions.

The objectives of this experiment were: a) to review the concept of simple acid-base reactions; b) to review the basic lab procedure of a titration and introduce the student to the concept of a primary standard and the process of standardization. From my graph my equivalence point was $pH = 8.80$, volume is 15 ml, $pK_a = 4.80$, $K_a = 1.58 \times 10^{-5}$ Introduction Acids dissolve in water with dissociation and the formation of hydrogen ions. According to the equation $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$. They are classed as strong or weak depending on the extent to which this dissociation occurs. Bases are substances that liberate hydroxide ions in water.

They too are classed as weak or strong. The Henderson-Hasselbach equation is derived from $pH = pK_a + \log \frac{base}{acid}$. In the experiment titration was used to study acid-base neutralization reaction quantitatively. In acid-base titration experiment, a solution of accurately known concentration was added gradually to another solution of known concentration until the chemical reaction between the two solutions were completed. The equivalence point was the point at which the acid was completely reacted with or neutralized by the base, and no HA remains; all is in the form of the conjugate base, A^- and water.

When these concentrations are equal $\log[A^-]/[HA]$ is zero and $pH = pK_a$. The point was signaled by a changing of color of an indicator that had been added to the acid solution. The common indicator was Phenolphthalein which was colorless in acidic and neutral solutions, but a pink was the result in basic solutions. This experiment has four objectives: (1) to standardize a NaOH solution, (2) to determine the molarity and pH of the acetic acid solution, (3) develop a titration curve to determine the molarity and K_a , (4) and perform a half-titration to determine the K_a of Acetic Acid. It was expected that the NaOH solution would have a 0.1 molarity, the experimental ionization constant of acetic acid would be close to $K_a = 1.76 \times 10^{-5}$.

Experimental The very first thing in beginning the experiment is to set up the apparatus, rinse out the apparatus using its chemicals that's going to be put in there, for example rinse out the burette using some HCl to get out any unwanted particles which may be present. Fill up the burette with HCl using a funnel. At this make sure there is protection in the eyes by wearing goggles so that nothing can get into the eyes. Whilst putting the HCl in the burette, make sure the burette is clamped into the stand firmly so to avoid any breakages.

When the burette is filled to a reasonable level, read how much HCl is in there, first make sure the funnel is out as this can alter the reading. When reading of the burette make sure to read below the meniscus at eye level. In the first part of the experiment standardization of the NaOH solution was done, in order to know the total concentration. A 500 ml of 0.1 M NaOH was diluted to an amount of 6.0 M NaOH. On an analytical balance 0.5 grams of KHP were measured using the bottle technique, and the total weight was

recorded. To standardize the supposed 0. M NaOH solution prepared in part 1, NaOH was titrated with potassium hydrogen phthalate of known amount. Complete neutralization, the reaction of one mole of NaOH with one mole of KHP, had occurred with the indicator phenolphthalein turning to a light pink color indicating that the titration has come to an endpoint. The volume was recorded. Three trials were performed for this part, and the molar concentration, and the average concentration were calculated. In the second part of the experiment, a weak acid (acetic acid) was titrated with a strong base NaOH in order to determine the average molarity, and the pH of the acetic acid solution.

First of all 100 ml of acetic acid solution of unknown molarity was given by our instructor. From there 10-30 ml of acetic acid solution was transferred to a beaker. A two point calibration with pH 4 and pH 7 buffer were to be performed in order to measure the pH of the acetic acid solution, and determine the $[H_3O^+]$ hydronium ions present. To do the titration the buret was filled with standardized hydroxide solution. The 25.00 ml sample of the acetic acid solution were to be transferred to a 125-ml Erlenmeyer flask, using 2-3 drops of phenolphthalein indicator.

The titration was performed three times. The K_a (acid ionization constant) was calculated using the pH and the average molarity of the acetic acid solution. In the third part of the experiment in order to determine the pH at half-neutralization for an unknown acid, we will measure the pH with a pH meter after each addition of base, and create a Titration curve. Before this was done we set up the reservoir, equilibrate it and fill it with standardized

NaOH(aq) solution. The tip of the reservoir was align so that it can be centered above the drop counter slot.

The drop counter was to be calibrated so that a precise volume of titrant was recorded in units of milliliters. A 10 ml graduated cylinder was place directly below the slot on the drop counter lining it up with the tip of the reagent reservoir. The top valve of the reservoir was open slowly so that drop by drop were released at a slow rate. When the volume of the NaOH reach between 9-10 ml the stopclock was close. The volume was recorded, and also the number of drops. First of all with the titration of a pH meter we refill the reservoir with the NaOH solution. Of the acetic acid solution 25. 0 ml were to be transfer to a 400 ml beaker. A magnetic stir bar was placed in the beaker containing the acetic acid solution, and also a pH sensor.

The titration was to be perform drop by drop until 50. 00 ml of the sodium hydroxide was dispensed, and the stopclock was close. The titration was repeated with a second sample of acetic acid to find the molar concentration of acetic acid, and determine the acid dissociation constant for acetic acid. In the fourth part of the experiment Half-Titration was perform in order to determine the K_a of the Acetic Acid solution. In this procedure 25. 0 ml Of 1. 00 M acetic acid was to be titrated to the endpoint with 1. 00 M sodium hydroxide. Then titrate it with the acetic acid until the solution is just slightly acidic just before the endpoint. In order to reach the endpoint again sodium hydroxide was added, and by doing this the pH was at a half-equivalence point. In a 250 ml beaker 25. 00 ml of 1. 00 M acetic solution was place, and then 2 drops of phenolphalein indicator were used.

The pH electrode was connected to read the pH, and two point calibration were performed with pH 4 and pH 7 buffers. The buret was filled with 1.0 M NaOH solution to 50 ml. From there the half-titration was performed. In order to begin this, the magnetic stirrer was placed in the beaker of the acetic acid solution. The stirring rate was to be gentle. In the buret the NaOH solution was added, and now the titrations were to be performed. When the titration reaches its equivalence point, the pH will increase rapidly and the indicator will change color. Thus in order to check the equivalence point, acetic acid was added to the solution so that it can be acidic, and the NaOH was added to reach to the equivalence point.