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CHM113

Determination of K_a : Titration of a Weak Acid

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Purpose

K_a , the acid dissociation constant, is a quantitative measure of the strength of an acid in solution. K_a is a constant for a given acid at a given temperature. The K_a (and pK_a) has a wide variety of applications, including water treatment systems, analysis of buffer solutions, and can be used in a general sense to measure the strength of an acid.

There are two ways to determine K_a by experiment, and this lab demonstrates them. One is to measure the pH of a solution containing a known concentration of weak acid. The other is to measure the pH at the half-neutralization point in the titration of weak acid with strong base. This allows one to determine K_a with a best fit line.

Procedure

The procedure was fairly straight forward, and essentially involved adding Sodium Hydroxide (NaOH), a base, drop-by-drop, to an acetic acid solution (CH_3COOH), while being stirred, and monitoring the pH changes.

We added 20 [mL] of the acetic acid to a 100 [mL] beaker on a stirrer. We then purged and cleaned 60 [mL] reaction reservoir, which was filled with 0.100 M NaOH solution, and set to drop at a rate of roughly 1 drop per second. The drops fell through a drop counter, which was connected to the computer, and configured to take a pH reading of the solution in the 100 [mL] beaker each time there was a drop. The beaker was stirred continuously at a moderate rate.

The drop counter was calibrated with NaOH before the experiment, allowing for an accurate calculation of the volume and ratio of the solution in the beaker.

Equations

Calibration of the Drop Counter

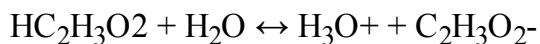
It was necessary to know the volume of a drop to determine the volume of the contents in the beaker. We found that

$$107 \text{ [drops]} / 5.20 \text{ [mL]} = 20.58 \text{ [drops / mL]} \text{ or } 0.48 \text{ [mL / drop]}$$

pH and pK_a (Henderson-Hasselbalch)

$$\text{pH} = \text{pK}_a + \log ([\text{A}^-]_{\text{eq}} / [\text{HA}]_{\text{eq}})$$

Reaction



(acid) (base) \leftrightarrow (conjugate acid) (conjugate base)

Data

We determined the equivalence point of the solution by examining graphs in Logger Pro. Based on the first derivative graph, the equivalence point was **40.52 [mL]**. The base equivalence point is the point when all the acid is titrated away; moles of acid = moles of base, and the base is 0.100 M.

The initial pH of the acid solution was **2.96**. The half equivalence point (20.26 mL) has a pH1/2 of **4.75**.

Molality of the Weak Acid

Molality of the weak acid was calculated to be **0.102 [M]**

$$\begin{aligned}M_{\text{acid}} &= [\text{base}] (V_{\text{base}} / V_{\text{acid}}) \\(0.100 [\text{M}]) (20.52 [\text{mL}] / 20.00 [\text{mL}]) &= \mathbf{0.102 [\text{M}]} \\K_a &= 1.09 \times 10^{-3} / 0.102[\text{M}] = 5.88 \times 10^{-6}\end{aligned}$$

K_a from pH of original solution.

The K_a can be calculated using the Henderson-Hasselbalch equation;

$$\begin{aligned}\text{pH} &= \text{pK}_a + \log ([\text{A}^-] / [\text{HA}]) \\[\text{H}^+] &= 10^{-2.96} = 1.09 \times 10^{-3}.\end{aligned}$$

There is a one to one molar ratio of [H⁺] to [A⁻]. Therefore,

$$\begin{aligned}[\text{A}^-] &= [\text{H}^+] = 1.09 \times 10^{-3}. \\K_a &= [(1.09 \times 10^{-3}) (1.09 \times 10^{-3})] / 0.102 [\text{M}] \\K_a &= 1.2 \times 10^{-5}\end{aligned}$$

K_a from pH of Half Neutralization Point.

The pH1/2 is **4.75**.

$$\text{pH}_{1/2} = \text{pK}_a = -\log K_a = 4.75$$

$$K_a = 10^{-4.75} = 1.7 * 10^{-5}$$

The K_a is $1.7 * 10^{-5}$.

Discussion

The lab process and calculations were fairly straightforward. We also found we had relatively accurate data (book value for the pK_a of acetic acid is 4.75, we found 4.76).

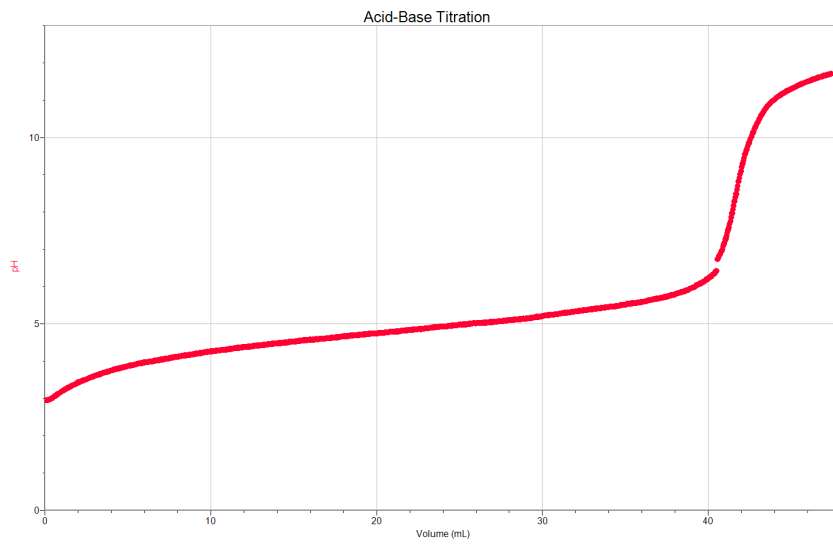
We had a small issue at one point. The lab manual instructs to lower the drip rate as the equivalence point is approached. When I attempted to adjust the rate, I temporarily stopped the dripping, resulting in no data points being recorded, while stirring (and mixing) still occurred. This resulted in a jump in pH when the next data point was taken after this incident. There was a jump of pH from 6.414 to 6.733.

Conclusions

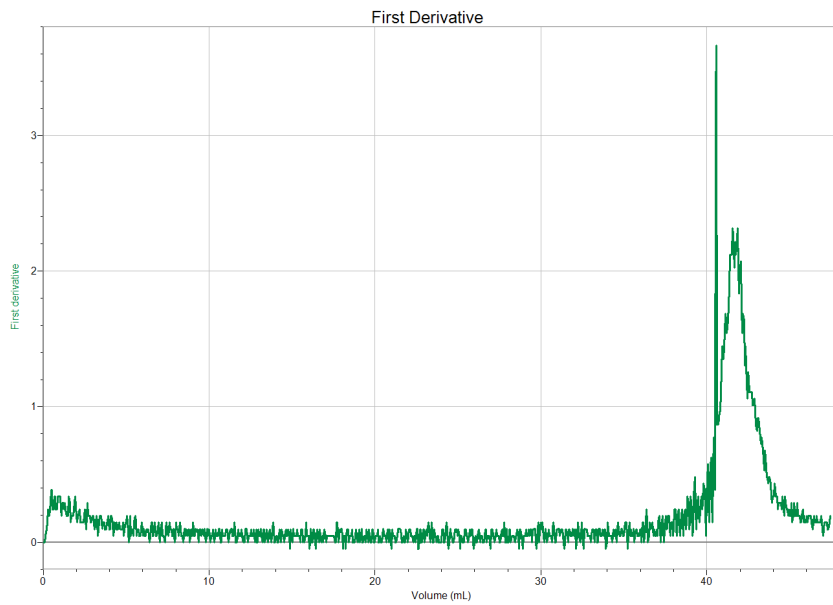
The lab was fairly straightforward, and demonstrated that it was relatively easy to determine the K_a of an acid, and use the K_a to calculate the pH of a given acid-base solution, using the Henderson-Hasselbalch equation.

Appendix

Titration Graph



First Derivative



Second Derivative (not used)

