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# *Lab 4: pKa of an Indicator*

## Abstract

pKa is the negative base-10 logarithm of the acid dissociation constant (Ka) of a solution. This value is often used to demonstrate the acid's strength; the lower it is, the stronger the acid is. To attain this value, procedures like visible spectroscopy, Beer's Law, Henderson Hasselbalch equation, and a pH meter. Using these procedures, a graph of Beer's plot was created to determine the slope, which was then used to determine the relationship between the concentration of the acid form of the indicator and its absorbance. A Henderson Hasselbalch equation was then used to create to determine the pKa, which was 5.8.

## Experimental

For this lab, the procedures were split into five separate potions. In the first procedure, the pKa was determined using visible spectroscopy, Beer's Law, and the Henderson Hasselbalch equation. Two test tubes were used. In the first test tube, 1mL of $8.0 * 10^{-5}$ M stock indicator and 5mL of 0.1M HCl were mixed. In the second test tube, 1mL of $8.0 * 10^{-5}$ M stock indicator and 5mL of 0.1M NaOH were mixed. A cuvette was then filled with each of these solutions and placed in the spectrophotometer. Their color and absorbance were recorded in *Results Table 1*.

In the second procedure, a Beer's law plot was created using various concentrations. To find these concentrations, five scintillation vials were each filled with 5mL of HCl. Then the first vial had 5mL of stock solution added and mixed. Once mixed, 5mL of Vial 1 soln was placed into Vial 2, then mixed. Next, 5mL of Vial 2 solution was placed in Vial 3 and mixed. This process was repeated for Vials 3 - 5.  Lastly, *Results Graph 1* was created with the data found.

In the third procedure, a rough estimate of the kPa was found using a pH meter. To start, about 100mL of distilled water and approximately 3mL of stock indicator were placed in a 250mL beaker. This beaker was then placed on a magnetic stir pad. Next, HCl was added drop by drop until the solution began to change color. Once the bright pink color was achieved, the drops were stopped. The pH once that color was achieved was placed in *Results Table 2*. Next, the NaOH was added drop by drop until the solution began to change color. Once the color was a pale yellow, the dropping was stopped, and the pH was placed in *Results Table 2.* The average of these 12 points was taken as the average pH.

In the fourth procedure, an array of incremental buffers were created to determine the unknown kPa. To begin, a vial of the same pH was created. Then, 10mL of this solution was placed into a separate test tube. The beaker then had its pH lowered by 0.2 by adding a few drops of the acid solution. Then, another 10mL of this solution was added to another beaker. This process was repeated until there were a total of seven buffer test tubes. The pH values for each test tube are in C1 for *Results Table 3.*

In the fifth and final procedure, a Henderson-Hasselbalch plot was created. In this plot, C1 = actual pH, C2 = the Absorbance, C3 = Indicator concentration, C4 = C2 / slope from Beer's Law plot, C5 = C3 - C4, C6 = C5 / C4, and C7 = log(C6). These values were placed in *Results Table 3*, then plotted to *Results Graph 2*.

**Results and Discussion**

*Table 1: The observed color and absorbance of the acid & base forms of the indicator*

| | *Observed Color* | *λ max (nm)* | *Absorbance* |
|---|---|---|---|
| *Acid form [HIn$^-$]* | Bright pink | 522 nm | 0.61 |
| *Base form [In$^-$]* | Pale yellow | 440 nm | 0.13 |

*Graph 1: Beer's Law plot.*

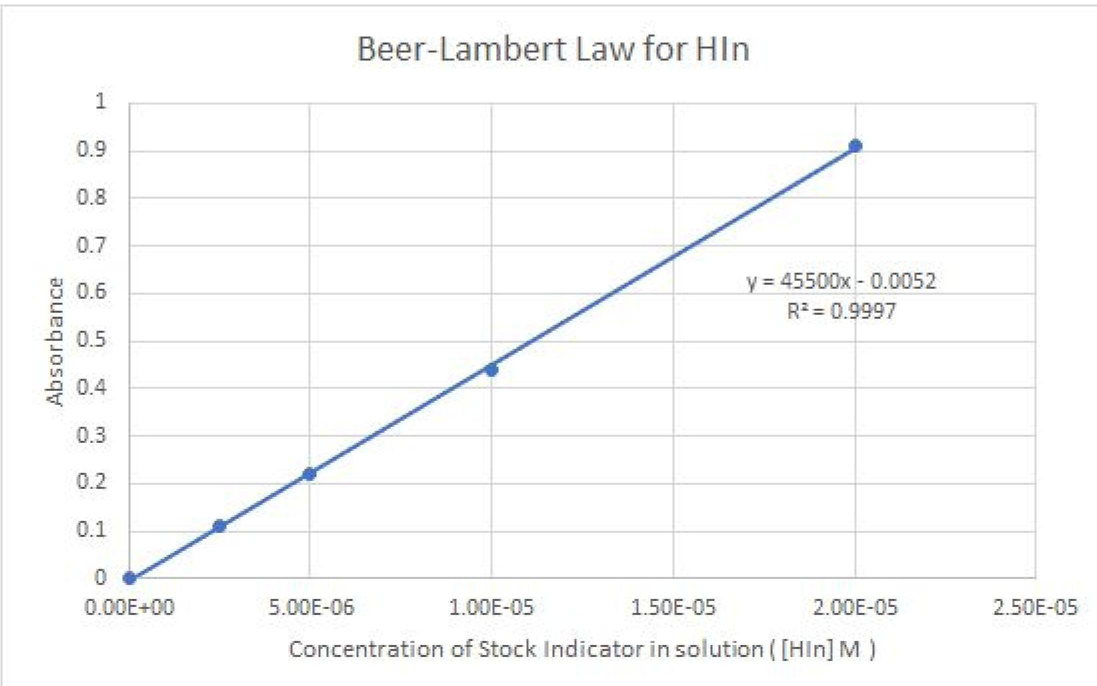



*Table 2: – pH data for the estimation of pKa*

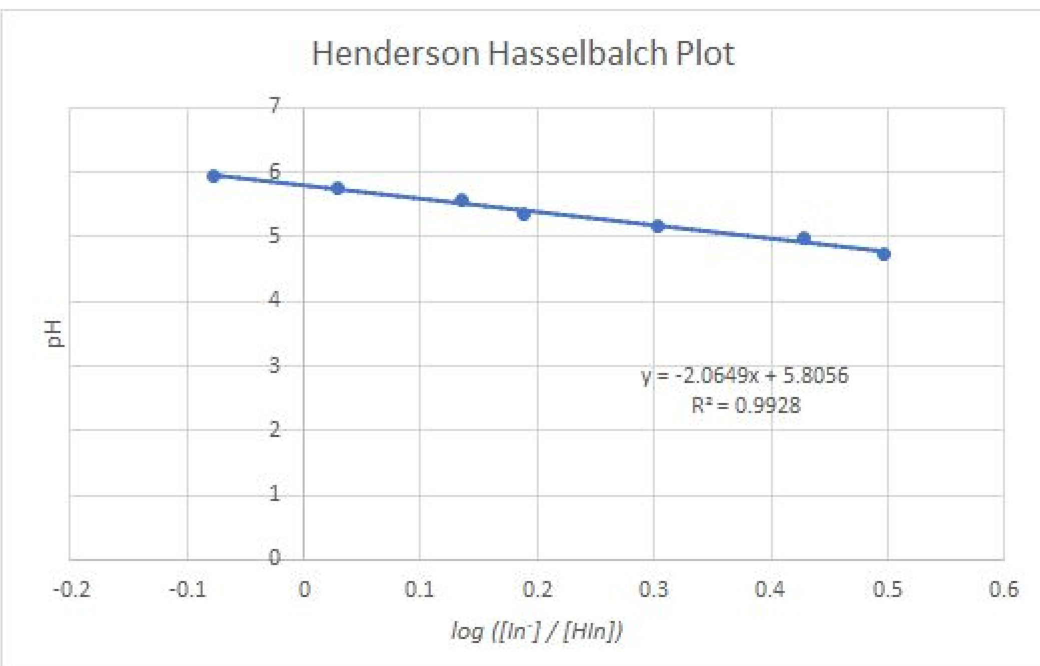
| *Vial Number* | *pH @ acid color* | *pH @ base color* |
|---|---|---|
| *1* | 3.94 | 7.43 |
| *2* | 3.63 | 8.73 |
| *3* | 3.75 | 8.63 |
| *4* | 3.83 | 8.96 |
| *5* | 3.77 | 7.97 |

| | | |
|---|---|---|
| *6* | 3.69 | 7.62 |
| *Average pH* | = **5.971** | = Estimated pKa |

*Table 3: Henderson- Hasselbalch Plot Data*

| | *C1* | *C2* | *C3* | *C4* | *C5* | *C6* | *C7* |
|---|---|---|---|---|---|---|---|
| *Test tube* | *Actual pH* | *Abs* | *[HIn]* | *[HIn]* | *[In⁻]* | *[In⁻] / [HIn]* | *Log (C6)* |
| *1* | 4.75 | 0.08 | 7.27 E-6 | 1.76 E-6 | 5.51 E-6 | 3.13 | 0.496212 |
| *2* | 4.97 | 0.09 | 7.27 E-6 | 1.98 E-6 | 5.2 9E-6 | 2.68 | 0.427387 |
| *3* | 5.18 | 0.11 | 7.27 E-6 | 2.42 E-6 | 4.85 E-6 | 2.01 | 0.302577 |
| *4* | 5.36 | 0.13 | 7.27 E-6 | 2.86 E-6 | 4.41 E-6 | 1.54 | 0.188788 |
| *5* | 5.56 | 0.14 | 7.27 E-6 | 3.08 E-6 | 4.19 E-6 | 1.36 | 0.134416 |
| *6* | 5.77 | 0.16 | 7.27E-6 | 3.52 E-6 | 3.75 E-6 | 1.07 | 0.02833 |
| *7* | 5.95 | 0.18 | 7.27E-6 | 3.96 E-6 | 3.31 E-6 | 0.84 | -0.07691 |

*Graph 2: Henderson Hasselbalch Graph using Actual pH and log ([In⁻] / [HIn])*



The trend for the plot is a decreasing function. The plot's y-intercept is the pKa, which is **5.8056,** and the average $\lambda$ max is 519.2. This means that the calculated pH of the indicator is the pKa value; however, the pKa should be 4.95. Possible sources of errors would include in Procedure 3. The drops of HCl and NaOH could have been more precise. The number of drops to make the solution acidic / basic when transitioning the pKa from acid to base form varied during each iteration of the experiment. To improve in the future, more time and precautions should be taken between each phase to give adequate drops.