

Determination of the Dissociation Constant of Weak Acids

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1 Data

Solution	pH
#1	5.31
#2	5.40

2 Calculations

We know that the K_a value is related to the concentrations of the major species in the solution as (for KHP where P is the phalate in Potassium hydrogen phalate)

$$K_a = \frac{[\text{H}_3\text{O}^+][P^{2-}]}{[\text{HP}^-]} \quad (1)$$

And, since we painstakingly made sure the concentrations had $[P^{2-}] = [\text{HP}^-]$, these cancel out leaving

$$K_a = [\text{H}_3\text{O}^+] \quad (2)$$

Since $\text{p}K_a = -\log K_a$ and we have shown that K_a is the hydronium concentration, $\text{pH} = -\log K_a$ which can be rewritten as

$$K_a = \frac{1}{10^{\text{pH}}} \quad (3)$$

Using this equation with the data, we can see that for the first trial,

$$K_a = \frac{1}{10^{5.31}} = 4.90 \times 10^{-6}$$

and the second,

$$K_a = \frac{1}{10^{5.40}} = 3.98 \times 10^{-6}$$

Taking the mean, we can say that for the acid, the $K_a = 4.44 \times 10^{-6}$

3 Discussion

- (a) $\text{NaHSO}_4 \rightarrow \text{Na}^+ + \text{HSO}_4^-$
 - (b) $\text{HSO}_4^- + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{SO}_4^{2-}$
 - (c) $K_a = \frac{[\text{H}_3\text{O}^+][\text{SO}_4^{2-}]}{[\text{HSO}_4^-]}$
 - (d) If we were to create a solution that we can guarantee has the equal concentrations of SO_4^{2-} and HSO_4^- by creating a solution of HSO_4^- from the salt, dividing it into two equal volumes, neutralizing one of the parts, and recombining it, we can then determine its pH to then feed into (3) because the concentrations would cancel out as shown in the calculation section.
- It is not necessary to know the exact mass of the acid whose K_a is to be determined because the solution is made with exact amounts of SO_4^{2-} and HSO_4^- which cancels in the K_a expression, so the actual amount of acid is irrelevant.
- The exact concentration of NaOH is not needed because we are finding the equivalence point for the solution to neutralize it. We do not need to know how much NaOH we are even adding. At the end, as long as we can be sure the solution is neutralized, the concentrations for the K_a expression will cancel out, making the amount or concentration of the neutralizer irrelevant.
- It is necessary to precisely measure the volume of the distilled water used to dissolve the acid so that we can precisely divide the solution into two solutions. If we can not precisely do this, we can not be sure we are neutralizing half of the solution, and we can not cancel out the concentrations with a clean conscience.
- The Henderson-Hasselbalch equation is

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]} \quad (4)$$

When the concentrations of the conjugate acid and base are equal, the $\frac{[\text{A}^-]}{[\text{HA}]}$ becomes 1 due to cancellation. Then, the $\log 1 = 0$ so the equation simplifies to

$$\text{pH} = \text{p}K_a \quad (5)$$

as was used for this lab.