Chapter 9 Sapling

1.

Given a diprotic acid, H_2A, with two ionization constants of $K_{a1} = 2.96\times10^{-4}$ and $K_{a2} = 3.40\times10^{-11}$, calculate the pH and molar concentrations of H_2A, HA^-, and A^{2-} for each of the solutions below.

(a) a 0.134 M solution of H_2A

pH	$[H_2A]$	$[HA^-]$	$[A^{2-}]$
Number 2.21	Number .128 M	Number 0.00615 M	Number 3.4×10^{-11} M

(b) a 0.134 M solution of NaHA

pH	$[H_2A]$	$[HA^-]$	$[A^{2-}]$
Number 7.00	Number 0.0000454 M	Number .134 M	Number 0.0000454 M

(c) a 0.134 M solution of Na_2A

pH	$[H_2A]$	$[HA^-]$	$[A^{2-}]$
Number 11.8	Number 3.38×10^{-11} M	Number 0.00613 M	Number .128 M

For part (a) treat H_2A as a monoprotic weak acid with $K_a = K_{a1}$.

$$\underset{F-x}{H_2A} \rightleftharpoons \underset{x}{H^+} + \underset{x}{HA^-} \qquad K_{a1} = \frac{x^2}{F-x} \qquad x^2 + K_{a1}x - K_{a1}F = 0$$

$$x = \frac{-K_{a1} \pm \sqrt{K_{a1}^2 + 4K_{a1}F}}{2} = \frac{-0.000296 + \sqrt{(0.000296)^2 + 4(0.000296)(0.134)}}{2}$$

$$= \frac{-0.000296 + 0.0126}{2} = 0.00615\ \text{M} = [H^+] = [HA^-] \qquad [H_2A] = 0.134 - 0.00615 = 0.128\ \text{M}$$

$$\text{pH} = -\log(0.00615) = 2.21 \qquad [A^{2-}] = \frac{[HA^-]}{[H^+]} \times K_{a2} = \frac{0.00615}{0.00615} \times 3.40\times10^{-11} = 3.40\times10^{-11}\ \text{M}$$

For part (b) note that HA^- is the intermediate form of a diprotic acid. It is an amphiprotic species that can both accept and donate a proton. The following equation is for calculating the H^+ concentration, where you assume $[HA^-] \approx F$, the formal concentration.

$$[H^+] = \sqrt{\frac{K_{a1}K_{a2}[HA^-] + K_{a1}K_w}{K_{a1} + [HA^-]}}$$