Answer on Question # 82831, Chemistry/General Chemistry

HA is a weak acid. Its ionization constant, Ka, is 4.3×10-13. Calculate the pH of an aqueous solution where the initial concentration of NaA is 0.047 M

Solution

$HA \leftrightarrow H^+ + A^-$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

As initial concentration of NaA, conjugate base, is given, then:

$A^- + H_2O \leftrightarrow HA + OH^-$

$$K_b = \frac{[HA][OH^-]}{[A^-]}$$

As we know Ka we can find Kb

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{4.3 \times 10^{-13}} = 2.326 \times 10^{-2}$$

We should use ICE table to find equilibrium concentrations of all species:

	A^-	HA	OH^-
Initial	0.047 M	0	0 M
Change	-x	+x	+x
Equilibrium	0.047 - x	x	x

$$K_b = \frac{[HA][OH^-]}{[A^-]}$$

$$2.326 \times 10^{-2} = \frac{(x)(x)}{(0.047 - x)}$$

Solve the quadratic equation:

$$x^2 = 2.326 \times 10^{-2}(0.047 - x)$$

$$x^2 + 2.326 \times 10^{-2} - 1.093 \times 10^{-3}$$

$$D = (2.326 \times 10^{-2})^2 - 4 \times (-1.093 \times 10^{-3}) = 4.913 \times 10^{-3}$$

$$x_{1,2} = \frac{-2.326 \times 10^{-2} \pm \sqrt{4.913 \times 10^{-3}}}{2}$$

$$x_1 = 2.342 \times 10^{-2}, \quad x_2 = -4.668 \times 10^{-2}$$

As value of concentration is positive number, then x =2.342×10^{-2}

$[OH^-] = 2.342 \times 10^{-2}$

As $[OH^-][H^+] = K_w$, then $[H^+] = Kw/[OH^-] = 1\times10^{-14}/2.342\times10^{-2} = 4.270\times10^{-13}$

$pH = -\log(4.270\times10^{-13}) = 12.37$

Answer: 12.37