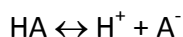


Answer on Question # 82831, Chemistry/General Chemistry

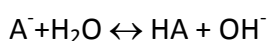
HA is a weak acid. Its ionization constant, K_a , is 4.3×10^{-13} . Calculate the pH of an aqueous solution where the initial concentration of NaA is 0.047 M

Solution



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

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



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

As we know K_a we can find K_b

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

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 $x = 2.342 \times 10^{-2}$

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

$$\text{As } [\text{OH}^-][\text{H}^+] = K_w, \text{ then } [\text{H}^+] = K_w / [\text{OH}^-] = 1 \times 10^{-14} / 2.342 \times 10^{-2} = 4.270 \times 10^{-13}$$

$$\text{pH} = -\log(4.270 \times 10^{-13}) = 12.37$$

Answer: 12.37

Answer provided by www.AssignmentExpert.com