Chapter 6
Equilibrium

Question 6.61:

The ionization constant of nitrous acid is 4.5×10^{-4}. Calculate the pH of 0.04 M sodiumnitrite solution and also its degree of hydrolysis.

Answer 6.61:

$NaNO_2$ is the salt of a strong base (NaOH) and a weak acid (HNO_2).

$$NO_2^- + H_2O \rightleftharpoons HNO_2 + OH^-$$

$$K_h = \frac{[HNO_2][OH^-]}{[NO_2^-]} \Rightarrow \frac{K_w}{K_a} = \frac{10^{-14}}{4.5 \times 10^{-4}} = 0.22 \times 10^{-10}$$

Now, If x moles of the salt undergo hydrolysis, then the concentration of various species present in the solution will be:

$[NO_2^-] = .04 - x$

$[HNO_2] = x$

$[OH^-] = x$

$$K_h = \frac{x^2}{0.04} = 0.22 \times 10^{-10} \Rightarrow x^2 = .0088 \times 10^{-10} \quad x = .093 \times 10^{-5}$$

$$\therefore [OH^-] = 0.093 \times 10^{-5} M \Rightarrow [H_3O^+] = \frac{10^{-14}}{.093 \times 10^{-5}} = 10.75 \times 10^{-9} M$$

$$\Rightarrow pH = -\log(10.75 \times 10^{-9}) = 7.96 \Rightarrow \frac{x}{0.04} = \frac{.093 \times 10^{-5}}{.04} = 2.325 \times 10^{-5}$$

Question 6.62:

A 0.02M solution of pyridinium hydrochloride has pH = 3.44. Calculate the ionization constant of pyridine.

Answer 6.62:

pH = 3.44

We know that,

pH = − log [H⁺]

$\therefore [H^+] = 3.63 \times 10^{-4}$

Then $K_h = \frac{(3.63 \times 10^{-4})^2}{0.02}$ (∵ Concentration = 0.002 M)

$\Rightarrow K_h = 6.6 \times 10^{-6}$

Now, $K_h = \frac{K_w}{K_a} \Rightarrow K_a = \frac{K_w}{K_h} = \frac{10^{-14}}{6.6 \times 10^{-6}} = 1.51 \times 10^{-9}$

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