

## Chapter 6

### Equilibrium

#### Question 6.61:

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

#### Answer 6.61:

$\text{NaNO}_2$  is the salt of a strong base ( $\text{NaOH}$ ) and a weak acid ( $\text{HNO}_2$ ).



$$K_b = \frac{[\text{HNO}_2][\text{OH}^-]}{[\text{NO}_2^-]} \Rightarrow \frac{K_b}{K_a} = \frac{10^{-14}}{4.5 \times 10^{-4}} = 0.22 \times 10^{-10}$$

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

$$[\text{NO}_2^-] = .04 - x$$

$$[\text{HNO}_2] = x$$

$$[\text{OH}^-] = x$$

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

$$\therefore [\text{OH}^-] = 0.093 \times 10^{-5} \text{ M} \Rightarrow [\text{H}_3\text{O}^+] = \frac{10^{-14}}{.093 \times 10^{-5}} = 10.75 \times 10^{-9} \text{ M}$$

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

#### Question 6.62:

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

#### Answer 6.62:

$$\text{pH} = 3.44$$

We know that,

$$\text{pH} = -\log [\text{H}^+]$$

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

$$\text{Then } K_b = \frac{(3.63 \times 10^{-4})^2}{0.02} \quad (\because \text{Concentration} = 0.002 \text{ M})$$

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

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

# Chemistry Chapter 6 Answers

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