



Question 41. The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it.

Answer:

$$\text{pH} = -\log [\text{H}^+] \text{ or } \log [\text{H}^+] = -\text{pH} = -3.76 = 4.24$$

$$\therefore [\text{H}^+] = \text{Antilog } 4.24 = 1.738 \times 10^{-4} = 1.74 \times 10^{-4} \text{ M}$$

Question 42. The ionization constant of HF, HCOOH and HCN at 298 K are  $6.8 \times 10^{-4}$ ,  $1.8 \times 10^{-4}$  and  $4.8 \times 10^{-9}$  respectively. Calculate the ionization constant of the corresponding conjugate base.

Answer:

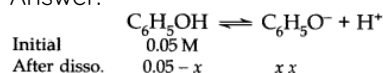
$$\text{For } \text{F}^-, K_b = K_w / K_a = 10^{-14} / (6.8 \times 10^{-4}) = 1.47 \times 10^{-11} = 1.5 \times 10^{-11}.$$

$$\text{For } \text{HCOO}^-, K_b = 10^{-14} / (1.8 \times 10^{-4}) = 5.6 \times 10^{-11}$$

$$\text{For } \text{CN}^-, K_b = 10^{-14} / (4.8 \times 10^{-9}) = 2.08 \times 10^{-6}$$

Question 43. The ionization constant of phenol is  $1.0 \times 10^{-10}$ . What is the concentration of phenolate ion in 0.05 M solution of phenol? What will be its degree of ionization if the solution is also 0.01 M in sodium phenolate?

Answer:



$$\therefore K_a = \frac{x \times x}{0.05 - x} = 1.0 \times 10^{-10} \text{ (Given) or } \frac{x^2}{0.05} = 1.0 \times 10^{-10}$$

$$\text{or } x^2 = 5 \times 10^{-12} \text{ or } x = 2.2 \times 10^{-6} \text{ M}$$

In presence of 0.01  $\text{C}_6\text{H}_5\text{ONa}$ , suppose  $y$  is the amount of phenol dissociated, then at equilibrium

$$\begin{array}{l} [\text{C}_6\text{H}_5\text{OH}] = 0.05 - y \approx 0.05, \\ [\text{C}_6\text{H}_5\text{O}^-] = 0.01 + y \approx 0.01 \text{ M}, [\text{H}^+] = y \text{ M} \end{array}$$

$$\therefore K_a = \frac{(0.01)(y)}{0.05} = 1.0 \times 10^{-10} \text{ (Given) or } y = 5 \times 10^{-10}$$

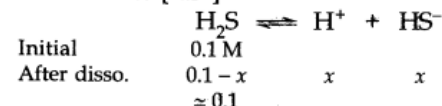
$$\therefore \alpha = \frac{y}{c} = \frac{5 \times 10^{-10}}{5 \times 10^{-2}} = 10^{-8}.$$

Question 44. The first ionization constant of  $\text{H}_2\text{S}$  is  $9.1 \times 10^{-8}$ .

Calculate the concentration of  $\text{HS}^-$  ions in its 0.1 M solution and how will this concentration be affected if the solution is 0.1 M in HCl also? If the second dissociation constant of  $\text{H}_2\text{S}$  is  $1.2 \times 10^{-13}$ , calculate the concentration of  $\text{S}^{2-}$  under both conditions.

Answer:

To calculate  $[\text{HS}^-]$

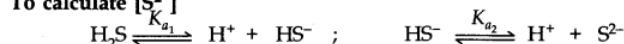


$$K_a = \frac{x \times x}{0.1} = 9.1 \times 10^{-8} \text{ or } x^2 = 9.1 \times 10^{-9} \text{ or } x = 9.54 \times 10^{-5}.$$

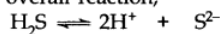
In presence of 0.1 M HCl, suppose  $\text{H}_2\text{S}$  dissociated is  $y$ . Then at equilibrium,  
 $[\text{H}_2\text{S}] = 0.1 - y \approx 0.1$ ,  $[\text{H}^+] = 0.1 + y \approx 0.1$ ,  $[\text{HS}^-] = y$  M

$$K_a = \frac{0.1 \times y}{0.1} = 9.1 \times 10^{-8} \text{ (Given)} \quad \text{or} \quad y = 9.1 \times 10^{-8} \text{ M}$$

To calculate  $[\text{S}^{2-}]$



For the overall reaction,



$$K_a = K_{a1} \times K_{a2} = 9.1 \times 10^{-8} \times 1.2 \times 10^{-13} = 1.092 \times 10^{-20}$$

$$K_a = \frac{[\text{H}^+]^2 [\text{S}^{2-}]}{[\text{H}_2\text{S}]}$$

In the absence of 0.1 M HCl,  $[\text{H}^+] = 2 [\text{S}^{2-}]$

Hence, if  $[\text{S}^{2-}] = x$ ,  $[\text{H}^+] = 2x$

$$\therefore \frac{(2x)^2 x}{0.1} = 1.092 \times 10^{-20} \quad \text{or} \quad 4x^3 = 1.092 \times 10^{-21} = 273 \times 10^{-24}$$

$$3 \log x = \log 273 - 24 = 2.4362 - 24$$

$$\log x = 0.8127 - 8 = \bar{8}.8127,$$

or

$$x = \text{Antilog } \bar{8}.8127 = 273 \times 10^{-24} = 6.497 \times 10^{-6} = 6.5 \times 10^{-6} \text{ M.}$$

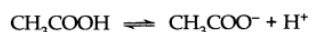
In presence of 0.1 M HCl, suppose  $[\text{S}^{2-}] = y$ , then

$$[\text{H}_2\text{S}] = 0.1 - y \approx 0.1 \text{ M}, \quad [\text{H}^+] = 0.1 + y \approx 0.1 \text{ M}$$

$$K_a = \frac{(0.1)^2 \times y}{0.1} = 1.09 \times 10^{-20} \quad \text{or} \quad y = 1.09 \times 10^{-19} \text{ M.}$$

Question 45. The ionization constant of acetic acid is  $1.74 \times 10^{-5}$ . Calculate the degree of dissociation of acetic acid in its 0.05 M solution. Calculate the concentration of acetate ions in the solution and its pH.

Answer:



$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} = \frac{[\text{H}^+]^2}{[\text{CH}_3\text{COOH}]}$$

$$\text{or} \quad [\text{H}^+] = \sqrt{K_a [\text{CH}_3\text{COOH}]} = \sqrt{(1.74 \times 10^{-5})(5 \times 10^{-2})} = 9.33 \times 10^{-4} \text{ M}$$

$$[\text{CH}_3\text{COO}^-] = [\text{H}^+] = 9.33 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log (9.33 \times 10^{-4}) = 4 - 0.9699 = 4 - 0.97 = 3.03$$

Question 46. It has been found that the pH of a 0.01 M solution of an organic acid is 4.15. Calculate the concentration of the anion, the ionization constant of the acid and its  $\text{pK}_a$ .

Answer:



$$\text{pH} = -\log [\text{H}^+] \quad \text{or} \quad \log [\text{H}^+] = -4.15 = \bar{5}.85$$

$\therefore$

$$[\text{H}^+] = 7.08 \times 10^{-5} \text{ M} = 7.08 \times 10^{-5} \text{ M}$$

$$[\text{A}^-] = [\text{H}^+] = 7.08 \times 10^{-5} \text{ M}$$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{(7.08 \times 10^{-5})(7.08 \times 10^{-5})}{10^{-2}} = 5.0 \times 10^{-7}$$

$$\text{pK}_a = -\log K_a = -\log (5.0 \times 10^{-7}) = 7 - 0.699 = 6.301$$

Question 47. Assuming complete dissociation, calculate the pH of the following solutions:

(a) 0.003 M HCl

(b) 0.005 M NaOH

(c) 0.002 M HBr

(d) 0.002 M KOH

Answer:

$$(a) \text{HCl} + \text{aq} \rightarrow \text{H}^+ + \text{Cl}^-, \therefore [\text{H}^+] = [\text{HCl}] = 3 \times 10^{-3} \text{ M}, \text{pH} = -\log (3 \times 10^{-3}) = 2.52$$

$$(b) \text{NaOH} + \text{aq} \rightarrow \text{Na}^+ + \text{OH}^-$$

$$\therefore [\text{OH}^-] = 5 \times 10^{-3} \text{ M}, [\text{H}^+] = 10^{-14} / (5 \times 10^{-3}) = 2 \times 10^{-12} \text{ M}$$

$$\text{pH} = -\log (2 \times 10^{-12}) = 11.70$$

$$(c) \text{HBr} + \text{aq} \rightarrow \text{H}^+ + \text{Br}^-, \therefore [\text{H}^+] = 2 \times 10^{-3} \text{ M}, \text{pH} = -\log (2 \times 10^{-3}) = 2.70$$

$$(d) \text{KOH} + \text{aq} \rightarrow \text{K}^+ + \text{OH}^-,$$

$$\therefore [\text{OH}^-] = 2 \times 10^{-3} \text{ M}, [\text{H}^+] = 10^{-14} / (2 \times 10^{-3}) = 5 \times 10^{-12}$$

$$\text{pH} = -\log (5 \times 10^{-12}) = 11.30$$

Question 48. Calculate the pH of the following solutions:

(a) 2g of TIOH dissolved in water to give 2 litre of the solution

(b) 0.3 g of  $\text{Ca}(\text{OH})_2$  dissolved in water to give 500 mL of the solution

(c) 0.3 g of NaOH dissolved in water to give 200 mL of the solution

(d) 1 mL of 13.6 M HCl is diluted with water to give 1 litre of the solution.

Answer:

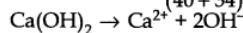
$$(a) \text{ Molar conc. of TlOH} = \frac{2g}{(204 + 16 + 1) g \text{ mol}^{-1}} \times \frac{1}{2 L} = 4.52 \times 10^{-3} M$$

$$[OH^-] = [TlOH] = 4.52 \times 10^{-3} M$$

$$[H^+] = 10^{-14} / (4.52 \times 10^{-3}) = 2.21 \times 10^{-12} M$$

$$\therefore pH = -\log (2.21 \times 10^{-12}) = 12 - (0.3424) = 11.66$$

$$(b) \text{ Molar conc. of Ca(OH)}_2 = \frac{0.3 g}{(40 + 34) g \text{ mol}^{-1}} \times \frac{1}{0.5 L} = 8.11 \times 10^{-3} M$$



$$[OH^-] = 2[Ca(OH)_2] = 2 \times (8.11 \times 10^{-3}) M = 16.22 \times 10^{-3} M$$

$$pOH = -\log (16.22 \times 10^{-3}) = 3 - 1.2101 = 1.79$$

$$pH = 14 - 1.79 = 12.21$$

$$(c) \text{ Molar conc. of NaOH} = \frac{0.3 g}{40 g \text{ mol}^{-1}} \times \frac{1}{0.2 L} = 3.75 \times 10^{-2} M$$

$$[OH^-] = 3.75 \times 10^{-2} M$$

$$pOH = -\log (3.75 \times 10^{-2}) = 2 - 0.0574 = 1.43$$

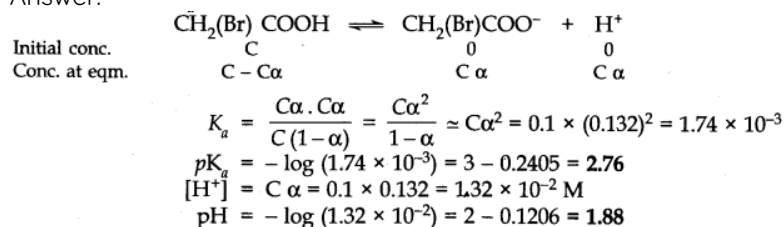
$$pH = 14 - 1.43 = 12.57$$

$$(d) M_1 V_1 = M_2 V_2 \therefore 13.6 M \times 1m L = M_2 \times 1000 mL \therefore M_2 = 1.36 \times 10^{-2} M$$

$$[H^+] = [HCl] = 1.36 \times 10^{-2} M, pH = -\log (1.36 \times 10^{-2}) = 2 - 0.1335 \approx 1.87$$

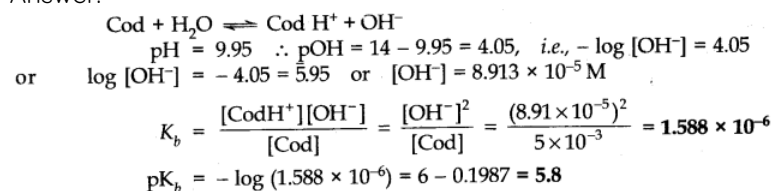
Question 49. The degree of ionization of a 0.1 M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the  $PK_a$  of bromoacetic acid.

Answer:



Question 50. The pH of 0.005 M codeine ( $C_{18}H_{21}NO_3$ ) solution is 9.95. Calculate the ionization constant and  $PK_b$ .

Answer:



\*\*\*\*\* END \*\*\*\*\*