

Daily Tutorial Sheet - 15

Level-3

- 159.** Millimoles of monobasic weak acid taken = Millimoles of NaOH required for end point
 $= 30 \times 0.1 = 3$

Upon addition of 20 ml alkali:

Number of millimoles of salt formed = 2

Number of millimoles of acid left = $3 - 2 = 1$

pH = 6.3

From Henderson's equation,

$$\text{pH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{Acid}]}$$

$$6.3 = \text{pK}_a + \log \frac{2}{1}$$

$$\therefore \text{pK}_a = 6.3 - 0.3 = 6 \text{ or } K_a = 10^{-6}$$

- 160.** The anion of the salt being conjugate base of a weak acid will hydrolyse in the following three steps.



K_{h_3} and K_{h_2} being quite smaller as compared to K_{h_1} , so consider only the 1st step hydrolysis so as to simplify the calculation.

The pOH of the solution may be given by the formula.

$$\text{pOH} = \frac{1}{2}[\text{pK}_w - \text{pK}_3 - \log C] = \frac{1}{2}[14 - 12 - \log 0.1] = \frac{1}{2} \times 3 = 1.5$$

So, $\text{pH} = 12.5$

- 161.** Let the number of millimole of weak base contained in 25 mL solution be x.

Upon addition of 10 mL 0.5 M HCl i.e. 5 millimole HCl, the salt formed will be 5 millimole and base remaining unreacted will be (x - 5) millimole.

For a base buffer mixture

$$[\text{OH}^-] = K_b \frac{[\text{base}]}{[\text{salt}]}$$

i.e. $\text{pOH} = \text{pK}_b + \log \frac{[\text{salt}]}{[\text{base}]}$

$\therefore \text{pH} = 8.6$

$\therefore \text{pOH} = 5.4 \text{ and } [\text{OH}^-] = 10^{-5.4}$

So, $10^{-5.4} = K_b \frac{[x - 5]}{5} \quad \dots\dots(1)$

Upon addition of 25 mL 0.5 M HCl i.e. 12.5 millimole HCl, the salt formed will be 12.5 millimole and base remaining unreacted will be (x - 12.5) millimole.

$$\therefore \text{pH} = 8$$

$$\therefore \text{pOH} = 6 \text{ and } [\text{OH}^-] = 10^{-6}$$

The mixture is again a basic buffer mixture for which

$$10^{-6} = \frac{K_b(x - 12.5)}{12.5} \quad \dots\dots (2)$$

Dividing equation (1) by (2)

$$10^{0.6} = \frac{2.5(x - 5)}{x - 12.5} \Rightarrow 4 = \frac{2.5(x - 5)}{x - 12.5} \Rightarrow 4x - 50 = 2.5x - 12.5$$

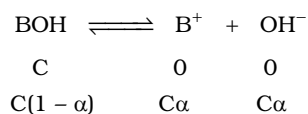
$$\therefore x = 25$$

$$\text{Molarity of base solution} = \frac{\text{No. of millimole}}{\text{Volume in mL}} = \frac{25}{25} = 1 \text{ M}$$

Putting the value of x in equation (2) :

$$K_b = 10^{-6}$$

Upon addition of 0 mL acid



$$[\text{OH}^-] = C\alpha = C\sqrt{\frac{K_b}{C}} = \sqrt{K_b \cdot C}$$

(neglecting α in comparison to unity which is true for concentrated solution).

$$= \sqrt{10^{-6} \times 1} = 10^{-3} \text{ M} \quad \therefore \text{pOH} = 3 \text{ and } \text{pH} = 11$$

At the end point :

The end point will reach when 50 mL 0.5 M HCl is added. 25 mL 1.0 M BOH i.e. 25 millimole of BOH will react with 50 × 0.5 HCl i.e. 25 millimole of HCl to form 25 millimole of salt. The total volume of solution will be 75 mL.

$$\therefore [\text{salt}] = \frac{25}{75} = \frac{1}{3} \text{ M}$$

The solution will be acidic due to salt hydrolysis. The pH of such a solution will be given by the expression.

$$\text{pH} = \frac{1}{2}[\text{pK}_a - \text{pK}_b - \log C] = \frac{1}{2}[14 - 6 - \log \frac{1}{3}] = 4.239$$

162. pH = 9.0

$$\therefore [\text{H}^+] = 10^{-9} \text{ M and } [\text{OH}^-] = 10^{-5} \text{ M}$$

$$[\text{Mg}^{2+}][\text{OH}^-] = 10^{-5} \text{ M}$$

$$[\text{Mg}^{2+}] = \frac{8.9 \times 10^{-12}}{(10^{-5})^2} = 8.9 \times 10^{-2} = 0.089 \text{ M}$$

Thus in saturated solution of $\text{Mg}(\text{OH})_2$, only 0.089 mol L^{-1} of Mg^{2+} will remain dissolved.

Hence the amount of Mg^{2+} to be precipitated out = $0.1 - 0.089 = 0.011 \text{ mol L}^{-1}$

- 163.** Molarity of ammonia may be calculated as,

$$M = \frac{x \times d \times 10}{M_B}$$

x = % solute by weight

d = density of solution

M_B = molecular weight of solute

$$M = \frac{10 \times 0.99 \times 10}{17} = 5.823$$

$$K_w = K_a \times K_b$$

$$10^{-14} = 5 \times 10^{-10} \times K_b$$

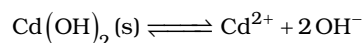
$$K_b = 2 \times 10^{-5}$$

We know, $[\text{OH}^-] = \sqrt{C \cdot K_b}$ for monoacidic base.

$$= \sqrt{5.823 \times 2 \times 10^{-5}} = 0.01079 \text{ M}$$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{10^{-14}}{0.01079} = 9.268 \times 10^{-12} \text{ M}$$

- 164.** The solubility equation is



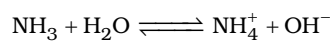
$$K_{sp} = [\text{Cd}^{2+}][\text{OH}^-]^2 = 2 \times 10^{-14}$$

Now $[\text{Cd}^{2+}] = 0.010 \text{ M}$ (coming from CdCl_2)

\therefore Maximum $[\text{OH}^-]$ so that precipitation does not occur is

$$[\text{OH}^-]^2 = \frac{2.0 \times 10^{-14}}{1.0 \times 10^{-2}} = 2 \times 10^{-12} \text{ or } [\text{OH}^-] = 1.4 \times 10^{-6} \text{ M}$$

Now $[\text{OH}^-]$ is derived from



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][1.4 \times 10^{-6}]}{[0.010]} \text{ or } [\text{NH}_4^+] = 1.3 \times 10^{-1} \text{ M}$$

\therefore NH_4^+ ion that must be supplied is 0.13 M

[Assuming very small dissociation of $\text{NH}_3 / \text{NH}_4\text{OH}$ in presence of NH_4Cl]