Ionic Equilibrium

Question1

At 27°C, 100 mL of 0.5 M HCl is mixed with 100 mL of 0.4 M NaOH solution. To this resultant solution, 800 mL of distilled water is added. What is the pH of final solution?

TG EAPCET 2024 (Online) 10th May Evening Shift

Options:

A. 12.0

B. 2.0

C. 1.3

D. 1.0

Answer: B

Solution:

To find the pH of the final solution, follow these steps:

Calculate the moles of HCl and NaOH before mixing:

HCl: Moles = $0.5\ \text{M} \times 0.1\ \text{L} = 0.05\ \text{mol}$

NaOH: Moles = $0.4\ \text{M} \times 0.1\ \text{L} = 0.04\ \text{mol}$

Determine the reaction outcome:

HCl and NaOH react in a 1:1 molar ratio as shown in the equation:

$\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$

Thus, 0.04 mol of HCl will completely react with 0.04 mol of NaOH. The remaining moles of HCl are:

$0.05\ \text{mol} - 0.04\ \text{mol} = 0.01\ \text{mol}$

Calculate the final volume after dilution:

By adding 800 mL of distilled water to 200 mL of the resultant solution, the total volume becomes:

$200\ \text{mL} + 800\ \text{mL} = 1000\ \text{mL} = 1\ \text{L}$

Determine the concentration of HCl after dilution:

$\text{Concentration of HCl} = \frac{0.01\ \text{mol}}{1\ \text{L}} = 0.01\ \text{M}$

Since HCl fully dissociates:

$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

Therefore, the concentration of H^+ ions is also 0.01 M.

Calculate the pH of the solution:

$\text{pH} = -\log[\text{H}^+]\text{pH} = -\log[0.01] = 2.0$

Thus, the pH of the final solution is 2.0.

Question2

A solution is prepared by mixing 10 mL of 1.0 M acetic acid and 20 mL of 0.5 M sodium acetate and diluted to 100 mL . If the $\text{p}K_a$ of acetic acid is 4.76, then the pH of the solution is

TS EAMCET 2023 (Online) 12th May Morning Shift

Options:

A. 4.76

B. 3.76

C. 5.76

D. 9.24

Answer: A

Solution:

The solution prepared by mixing the given quantities of acetic acid and sodium acetate is an acidic buffer.

To find the pH of an acidic buffer, we use the Henderson-Hasselbalch equation:

$\text{pH} = \text{p}K_a + \log \frac{[\text{Salt}\ ]}{[\text{Acid}\ ]}$

Given:

$\text{p}K_a = 4.76$ (provided)

Concentration of the salt, [Salt], is calculated as:

$$\frac{20\ \text{mL}\ \times 0.5\ \text{M}}{100\ \text{mL}} = 0.1\ \text{M}$$

Concentration of the acid, [Acid], is calculated as:

$$\frac{10\ \text{mL}\ \times 1.0\ \text{M}}{100\ \text{mL}} = 0.1\ \text{M}$$

Substituting these values into the Henderson-Hasselbalch equation:

$$\text{pH} = 4.76 + \log \frac{0.1}{0.1} = 4.76 + \log 1 = 4.76$$

Therefore, the pH of the solution is 4.76.
