

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:

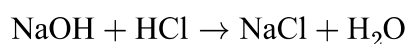
Calculate the moles of HCl and NaOH before mixing:

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:



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:



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 pK_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:

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



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.

