Marks
3

- A solution is prepared by dissolving 0.050 mol of acetic acid, 0.020 mol of sodium acetate and 0.0010 mol of HCl in water to give a final volume of 250 mL. The pK_a of acetic acid is 4.76. What is the pH of this solution?

HCl will react with the acetate to produce acetic acid:

$$\mathbf{HCl + CH_3CO_2^- \rightarrow CH_3COOH + Cl^-}$$

As 0.020 mol of $CH_3CO_2^-$ is initially present, the number of moles after this reaction is (0.020 – 0.0010) mol = 0.019 mol.

As 0.050 mol of CH_3COOH^- is initially present, the number of moles after this reaction is (0.050 + 0.0010) mol = 0.051 mol.

The final volume is 250 mL so the concentrations are:

$$[CH_3CO_2^-(aq)] = \frac{\text{number of moles}}{\text{volume}} = \frac{0.019\text{ mol}}{0.250\text{ L}} = 0.076\text{ M}$$

$$[CH_3COOH(aq)] = \frac{\text{number of moles}}{\text{volume}} = \frac{0.051\text{ mol}}{0.250\text{ L}} = 0.204\text{ M}$$

The pH of the solution containing both an acid (CH_3COOH) and its conjugate base ($CH_3CO_2^-$) is given by the Henderson-Haseelbalch equation:

$$\text{pH} = \text{p}K_a + \log\frac{[\text{base}]}{[\text{acid}]} = 4.76 + \log\frac{0.076}{0.204} = 4.33$$

pH = **4.33**