Marks 4

- What is the pH of a solution which is 0.10 M in both acetic acid and sodium acetate? The $K_a$ for acetic acid is $1.8 \times 10^{-5}$.

**By definition, p$K_a$ = -log$_{10}K_a$. Hence:**

$$pK_a = -\log(1.8 \times 10^{-5}) = 4.74$$

**Using the Henderson-Hasselbalch equation,** $pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$

**With [base] = [acetate] = 0.10 M and [acid] = [acetic acid] = 0.10 M, therefore:**

$$pH = 4.74 + \log \frac{0.10}{0.10} = 4.74$$

Answer: **4.74**

What is the final pH if 0.010 mol of HCl is added to 1.0 L of the above solution?

**Using number of moles = concentration × volume, the number of moles of acetate and acetic acid originally present in 1.0 L of a 0.10 M solutions are:**

$$n_{\text{acetate}} = 0.10 \text{ mol L}^{-1} \times 1.0 \text{ L} = 0.10 \text{ mol}$$
$$n_{\text{acetic acid}} = 0.10 \text{ mol L}^{-1} \times 1.0 \text{ L} = 0.10 \text{ mol}$$

**The added $H_3O^+$ from HCl will react with the acetate to produce more acetic acid:**

$$\text{acetate} + H_3O^+ \rightarrow \text{acetic acid} + H_2O$$

**Hence, after addition of HCl, the amount of acetate will decrease and the amount of acetic acid will increase:**

$$n_{\text{acetate}} = (0.10 - 0.010) \text{ mol} = 0.09 \text{ mol}$$
$$n_{\text{acetic acid}} = (0.10 + 0.010) \text{ mol} = 0.11 \text{ mol}$$

**Using concentration = number of moles / volume, their concentrations will become:**

$$[\text{acetate}] = 0.09 \text{ mol} / 1.0 \text{ L} = 0.09 \text{ M}$$
$$[\text{acetic acid}] = 0.11 \text{ mol} / 1.0 \text{ L} = 0.11 \text{ M}$$

**Using the Henderson-Hasselbalch equation,**

$$pH = 4.74 + \log \frac{0.09}{0.11} = 4.65$$

Answer: **4.65**