

- 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 1.8×10^{-5} .

By definition, $pK_a = -\log_{10}K_a$. 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 $[\text{base}] = [\text{acetate}] = 0.10 \text{ M}$ and $[\text{acid}] = [\text{acetic acid}] = 0.10 \text{ M}$, therefore:

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

Answer: 4.74

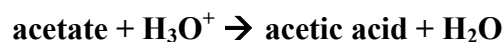
What is the final pH if 0.010 mol of HCl is added to 1.0 L of the above solution?

Using number of moles = concentration \times 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:



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