Marks 6

- Calculate the pH of a 0.200 M solution of acetic acid, $CH_3COOH$, at 25 ºC. (The p$K_a$ of acetic acid is 4.76).

**As acetic acid is a weak acid, $[H_3O^+]$ must be calculated:**

| | $CH_3COOH$ | $H_2O$ | $\rightleftharpoons$ | $H_3O^+$ | $CH_3COO^-$ |
|---|---|---|---|---|---|
| **initial** | **0.200** | **large** | | **0** | **0** |
| **change** | **$-x$** | **negligible** | | **$+x$** | **$+x$** |
| **final** | **$0.200 - x$** | **large** | | **$x$** | **$x$** |

**The equilibrium constant $K_a$ is given by:** $K_a = \dfrac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = \dfrac{x^2}{0.2 - x}$

**As p$K_a$ = 4.76 = $-\log_{10}K_a$ so $K_a = 10^{-4.76}$. As $K_a$ is very small, $0.200 - x \sim 0.200$ and hence:**

$x^2 = 0.200 \times 10^{-4.76}$ **or** $x = 0.0019$ **M** $= [H_3O^+]$

**Hence, the pH is given by:**

$\text{pH} = -\log_{10}[H_3O^+] = -\log_{10}[0.0019] = 2.73$

pH = **2.73**

Solid sodium acetate, $NaCH_3CO_2$, (0.15 mol) was dissolved in 0.500 L of 0.200 M acetic acid and the volume made up to 750 mL with water. What is the pH of the resulting solution?

**The solution contains a weak acid (acetic acid) and its conjugate base (acetate). 0.15 mol of acetate is present in 750 mL so its concentration is:**

**[base] = (0.15 mol) / (0.750 L) = 0.20 M**

**500 mL of 0.200 M acid contains (0.5 L) × (0.200 M) = 0.100 mol. The concentration of the acid in 750 mL is therefore:**

**[acid] = (0.100 mol) / (0.750 L) = 0.133 M**

**The Henderson-Hasselbalch equation can be used for this buffer:**

$$\text{pH} = \text{p}K_a + \log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right) = 4.76 + \log_{10}\left(\frac{0.20}{0.133}\right) = 4.94$$

pH = **4.94**

**ANSWER CONTINUES ON THE NEXT PAGE**

How much more $NaCH_3CO_2$ needs to be dissolved in the above solution to give a final pH of 5.00?

**A pH of 5.00 will be obtained when:**

$$\mathbf{pH = 4.76 + \log_{10}\left(\frac{[base]}{[acid]}\right) = 5.00 \text{ or } \log_{10}\left(\frac{[base]}{[acid]}\right) = 0.24}$$

**Hence,**

$$\mathbf{\left(\frac{[base]}{[acid]}\right) = 10^{0.24} = 1.74 \text{ or } [base] = 1.74 \times [acid] = 1.74 \times 0.133 = 0.232\ M}$$

**The number of moles of base in 750 mL is therefore (0.232 M) × (0.750 L) = 0.174 mol.**

**As 0.15 mol was added originally, an additional (0.17 – 0.15) = 0.02 mol is required.**

Answer: **0.02 mol**