Question#46462 – Chemistry – Inorganic Chemistry

Question:

Calculate the pH of 1.0 M solution of acetic acid. To what volume one liter of this solution be diluted so that the pH of the solution that is formed will be twice of original volume.[K_a = 1.8 ×10^{-5}]

Answer:

Acetic acid CH_3COOH is a weak acid and it dissociated in water solution to some extent according to equation:

$CH_3COOH \leftrightarrow CH_3COO^- + H^+$

Since the process is reversible, the constant of equilibrium of this process $K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$

The degree of dissociation for CH_3COOH is not great, than we can neglect the amount of CH_3COOH that was ionized comparing with the initial concentration. And, according to the reaction equation, the amount of H^+ and CH_3COO^- formed are the same. Using this consideration:

$$K_a = \frac{[H^+]^2}{[CH_3COOH]_0} \Rightarrow \sqrt{[H^+]} = K_a \times [CH_3COOH]_0$$

We have, that $[CH_3COOH]_0$ = 1.0 M and K_a = $1.8 \cdot 10^{-5}$. Therefore,

$$[H^+] = \sqrt{K_a \times [CH_3COOH]_0} = \sqrt{1.8 \cdot 10^{-5} \times 1.0} = 0.0042M$$

pH function is a negative logarithm from $[H^+]$:

$$pH = -\log[H^+] = -\log(0.0042) = 2.38$$

pH after the dilution has to be twice of original volume pH = 2.38×2 = 4.76. The corresponding H^+ concentration is:

$$pH = -\log[H^+] \Rightarrow [H^+] = 10^{-pH} = 10^{-4.76} = 1.74 \cdot 10^{-5} M$$

The corresponding initial concentration of CH_3COOH, which produced this amount of H^+:

$$[H^+] = \sqrt{K_a \times [CH_3COOH]_0} \Rightarrow [CH_3COOH]_0 = \frac{[H^+]^2}{K_a} = \frac{(1.74 \cdot 10^{-5})^2}{1.8 \cdot 10^{-5}} = 1.68 \cdot 10^{-5} M$$

The amount of moles remains constant, only volume of the solution has changed. If the initial volume V_1 = 1 L, hence:

$$C_1V_1 = C_2V_2 \Rightarrow V_2 = \frac{C_1V_1}{C_2} = \frac{1 \times 1.0}{1.68 \cdot 10^{-5}} \approx 6 \cdot 10^5 L$$