# Answer on Question #64937 – Chemistry – General Chemistry

Calculate the equilibrium concentrations of the dissociated species of acetic acid ($CH_3CHOOH$) in a 0.01 molar solution of acetic acid. Acetic acid is a weak acid its dissociation equilibrium constant at 25 degrees Celsius and 1 bar is $10^{-4.75}$. What is the pH equilibrium? What is degree of dissociation at equilibrium? %dissociation of an acid=100*(amount of acid dissociated/amount of acid originally present).

**Solution.**

$K_a$ ($CH_3COOH$)= $10^{-4.75}$ = $1.78\times10^{-5}$

$K_a$ = ([$H^+$] [$A^–$]) / [HA] = ([$H^+$] [$CH_3COO^–$]) / [$CH_3CHOOH$]

$K_a = x\cdot x / (0.01 - x) \approx x^2 / 0.01 = 1.78\times10^{-5}$

$x = \sqrt{1.78\times10^{-5}\times0.01} = 0.42\times10^{-3} = 4.2\times10^{-4}$ M

[$H^+$] = [$CH_3COO^–$] = $4.2\times10^{-4}$ M

[$CH_3CHOOH$] = 0.01 M

pH = $-\log$ [$H^+$] = $-\log (4.2\times10^{-4})$ = 3.38

$\alpha$ = [$H^+$] / [HA] = [$H^+$] / [$CH_3CHOOH$] × 100 % = $4.2\times10^{-4}$ / 0.01 × 100 % = 4.2 %

**Answer:** [$H^+$] = [$CH_3COO^–$] = $4.2\times10^{-4}$ M; [$CH_3CHOOH$] = 0.01 M

pH = 3.38

$\alpha$ = 4.2 %