

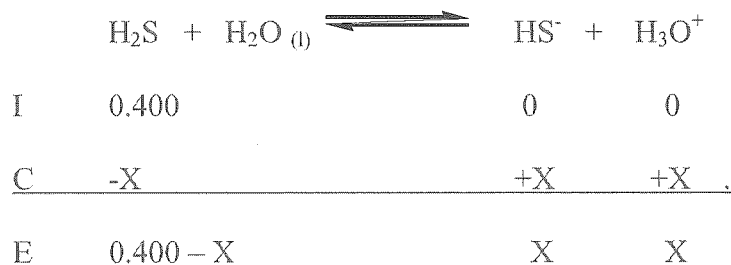
## 10. pH and pOH Calculations with Weak Acids and Bases

### a) Weak Acids

i) Example: What is the pH of 0.400 M H<sub>2</sub>S?

① Not 100% dissociated so [H<sub>2</sub>S] ≠ [H<sub>3</sub>O<sup>+</sup>]; so can't just pH = -log[0.400]

② The weak acid will form an equilibrium:



$$\textcircled{3} K_a = 9.1 \times 10^{-8} = \frac{[\text{HS}^-][\text{H}_3\text{O}^+]}{[\text{H}_2\text{S}]} = \frac{X^2}{0.400 - X}$$

④ Assume 0.400 - X = 0.400, because it is a weak acid and % dissociation < 5%

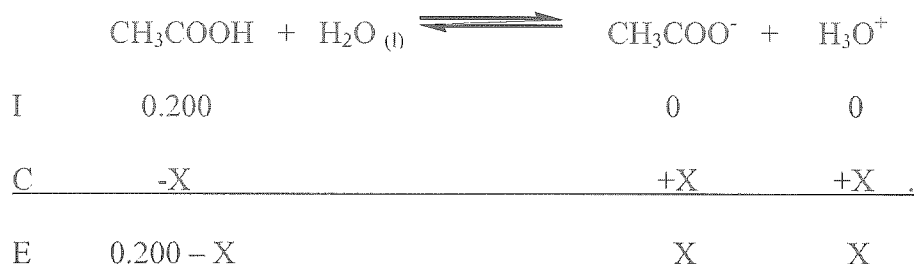
$$\textcircled{5} 9.1 \times 10^{-8} = \frac{X^2}{0.400} \quad X = 1.91 \times 10^{-4} = [\text{H}_3\text{O}^+]$$

$$\textcircled{6} \text{pH} = -\log[1.91 \times 10^{-4}] = 3.719$$

ii) Example: What is pH of 0.200 M CH<sub>3</sub>COOH?

① Not 100% dissociated so [CH<sub>3</sub>COOH] ≠ [H<sub>3</sub>O<sup>+</sup>]; so can't just pH = -log[0.200]

② The weak acid will form an equilibrium:



③  $K_a = 1.8 \times 10^{-5} = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = \frac{X^2}{0.200 - X}$

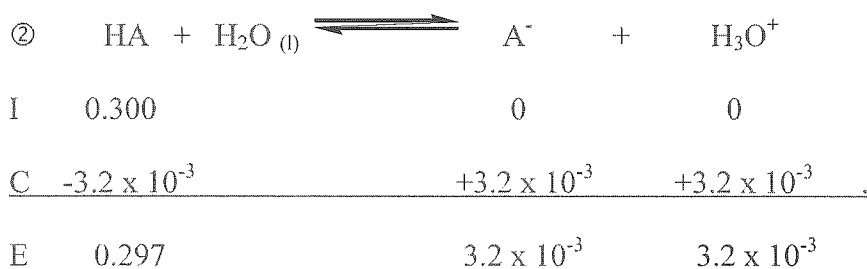
④ Assume 0.200 - X = 0.200, because it is a weak acid and % dissociation < 5%

⑤  $1.8 \times 10^{-5} = \frac{X^2}{0.200} \quad X = 1.90 \times 10^{-3} = [\text{H}_3\text{O}^+]$

⑥ pH = -log[1.90 x 10<sup>-3</sup>] = 2.721

iii) Example: What is K<sub>a</sub> for an unknown acid which gives a pH of 2.50 for a 0.300M solution?

① Equilibrium [H<sub>3</sub>O<sup>+</sup>] = antilog(-2.50) = 3.2 x 10<sup>-3</sup> M



③  $K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{(3.2 \times 10^{-3})^2}{0.297} = 3.4 \times 10^{-5}$

iv) Example: What is pH for a solution containing 0.300 M  $\text{NH}_4\text{NO}_3$ ?

①  $\text{NH}_4\text{NO}_3 \rightarrow \text{NH}_4^+ + \text{NO}_3^-$ ;  $\text{NO}_3^-$  is a spectator and  $\text{NH}_4^+$  acts as a weak acid



	I	0.300		0	0
	C	-X		+X	+X

	E	0.300 - X		X	X
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$$\text{③ } K_a = 5.6 \times 10^{-10} = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = \frac{X^2}{0.300 - X}$$

④ Assume  $0.300 - X = 0.300$ , because it is a weak acid and % dissociation < 5%

$$\text{⑤ } 5.6 \times 10^{-5} = \frac{X^2}{0.300} \quad X = 1.30 \times 10^{-5} = [\text{H}_3\text{O}^+]$$

$$\text{⑥ } \text{pH} = -\log[1.30 \times 10^{-5}] = 4.886$$

### b) Weak Bases

i) Example: What is the pOH and pH of 0.150 M  $\text{NH}_3$ ?

① Not 100% dissociated so  $[\text{NH}_3] \neq [\text{OH}^-]$



	I	0.150		0	0
	C	-X		+X	+X
	E	0.150 - X		X	X

$$\text{③ } \text{Must 1}^{\text{st}} \text{ find the } K_b. \quad K_b(\text{NH}_3) = \frac{K_w}{K_a(\text{NH}_4^+)} = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.78 \times 10^{-5}$$

$$\text{④ } K_b = \frac{[\text{OH}^-][\text{NH}_4^+]}{[\text{NH}_3]} = 1.78 \times 10^{-5} = \frac{(X)^2}{0.150 - X}$$

⑤ Assume  $0.150 - X = 0.150$

$$\text{⑥ } 1.78 \times 10^{-5} = \frac{X^2}{0.150} \quad X = 1.63 \times 10^{-3} \text{ M} = [\text{OH}^-]$$

$$\text{⑦ } \text{pOH} = -\log(1.63 \times 10^{-3}) = 2.788; \quad \text{pH} = 14 - 2.788 = 11.212$$

