Pre-AP Chemistry/AP Chemistry
Unit #12—Acids and Bases

Solving for pH of a Weak Base

Step 1: Write the Ionization Reaction and Identify Acid, Base, Conjugate Acid, and Conjugate Base.
Step 2: Write the Base Equilibrium Expression.
Step 3: Use ICE Chart
Step 4: Solve for "x" using the Base Equilibrium Expression.
Step 5: Solve for pH, pOH, $[H^+]$, and $[OH^-]$ at Equilibrium.
Step 6: Solve for Percent Ionization of the $[OH^-]$.

Write the ionization reaction, Ka and calculate pH, pOH, $[H^+]$, $[OH^-]$ and the percent ionization of a 0.350 M aqueous solution of methylamine where the $K_b = 4.38 \times 10^{-4}$.

Step 1: Write the Ionization Reaction and Identify Acid, Base, Conjugate Acid, and Conjugate Base

$$\underset{\text{Base}}{CH_3NH_2\ (aq)} + \underset{\text{Acid}}{H_2O\ (l)} \leftrightarrow \underset{\text{Conjugate Acid}}{CH_3NH_3^+\ (aq)} + \underset{\text{Conjugate Base}}{OH^-\ (aq)}$$

Step 2: Write the Acid Equilibrium Expression

$$K_b = \frac{[CH_3NH_3^+]\ [OH^-]}{[CH_3NH_2]}$$

Step 3: Use ICE Chart

	CH_3NH_2 (aq)	$CH_3NH_3^+$ (aq)	OH^- (aq)
Initial Concentration	0.35 M	0 M	0 M
Change in Concentration	~~- x M~~	+ x M	+ x M
Equilibrium Concentration	0.35 M	x M	x M

Step 4: Solve for "x" using the Acid Equilibrium Expression.

$$K_b = \frac{[x]\ \ [x]}{[0.35]} = 4.38 \times 10^{-4}$$

$$x^2 = 1.533 \times 10^{-4}\ M$$

$$x = 0.012\ M = [OH^-]_{eq} = [CH_3NH_3^+]_{eq}$$

Step 5: Solve for pH, pOH, $[H^+]$, and $[OH^-]$ at Equilibrium.

$$[H^+]_{eq} = 10^{-(12.093)} = 8.072 \times 10^{-13}\ M$$

$$pH = 14 - 1.907 = 12.093$$

$$pOH = -\log [0.012] = 1.907$$

$$[OH.]^{eq} = 0.012\ M$$

Step 6: Solve for Percent Ionization of the $[OH^-]$.

$$\frac{0.012\ M}{0.350\ M} \times 100 = 3.4\ \%\ OH^-$$

Solving for K_b of a Weak Acid

Step 1: Write the Ionization Reaction and Identify Acid, Base, Conjugate Acid, and Conjugate Base.
Step 2: Write the Base Equilibrium Expression.
Step 3: Find the concentration of the $[OH^-]_{eq}$ through pH
Step 4: Use ICE Chart
Step 5: Solve for K_b
Step 6: Solve for pH, pOH, $[H^+]$, and $[OH^-]$ at Equilibrium.
Step 7: Solve for Percent Ionization of the $[OH^-]$.

The pH of a 0.500 M solution of dimethylamine, $(CH_3)_2NH$, is 12.25. What is the Kb, pH, pOH, $[H^+]$, and $[OH^-]$ at equilibrium and percent ionization?

Step 1: Write the Ionization Reaction and Identify Acid, Base, Conjugate Acid, and Conjugate Base.

$$\underset{\text{Base}}{(CH_3)_2NH\ (aq)} + \underset{\text{Acid}}{H_2O\ (l)} \leftrightarrow \underset{\text{Conjugate Acid}}{(CH_3)_2NH_2^+\ (aq)} + \underset{\text{Conjugate Base}}{OH^-\ (aq)}$$

Step 2: Write the Base Equilibrium Expression.

$$K_b = \frac{[(CH_3)_2NH_2^+]\ \ [OH^-]}{[(CH_3)_2NH]}$$

Step 3: Find the concentration of the $[OH^-]_{eq}$ through pH

$$pOH = 14 - 12.25 = 1.75$$
$$[OH^-]_{eq} = 10^{-(1.75)} = 0.018\ M$$

Step 4: Use ICE Chart

	$(CH_3)_2NH\ (aq)$	$(CH_3)_2NH_2^+\ (aq)$	$OH^-\ (aq)$
Initial Concentration	0.100 M	0 M	0 M
Change in Concentration	- 0.018 M	+ 0.018 M	+ 0.018 M
Equilibrium Concentration	0.482 M	0.018 M	0.018 M

Step 5: Solve for K_b

$$K_b = \frac{[0.004]\ \ [0.004]}{[0.096]} = 6.722 \times 10^{-4}$$

Step 6: Solve for pH, pOH, $[H^+]$, and $[OH^-]$ at Equilibrium.

$$[H^+]_{eq} = 10^{-(12.25)} \qquad 5.623 \times 10^{-13}\ M$$

$$pH = 12.25 = 2.44$$

$$pOH = 14 - 2.44 = 1.75$$

$$[OH.]^{eq} = 0.018$$

Step 7: Solve for Percent Ionization of the $[OH^-]$.

$$\frac{0.018\ M}{0.100\ M} \times 100 = 3.6\ \%\ OH^-$$