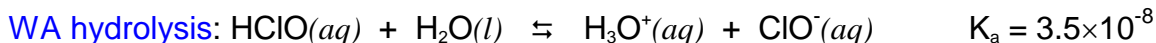


WEAK ACID, WEAK BASE AND SALT PROBLEMS KEY

1. a) Calculate the equilibrium concentrations of H_3O^+ , ClO^- and HClO in a 0.50 M solution of hypochlorous acid, $\text{HClO}(\text{aq})$. b) What is the pH of this solution? **HClO is WA!**



	$\text{HClO}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{ClO}^-(\text{aq})$			
I	0.50		0	0
C	-x		+x	+x
E	0.50 - x		x	x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{ClO}^-]}{[\text{HClO}]} \quad 3.5 \times 10^{-8} = \frac{x^2}{0.50 - x}$$

Assume x is much smaller than the initial [HClO] $\Rightarrow 0.50 - x = 0.50$

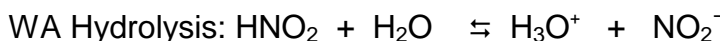
$$\frac{x^2}{0.50} = 3.5 \times 10^{-8} \quad x^2 = 0.50(3.5 \times 10^{-8}) = 1.75 \times 10^{-8} \quad x = \sqrt{1.75 \times 10^{-8}} = 1.32 \times 10^{-4}$$

- check assumption:** $\frac{x}{[\text{HA}]_{\text{initial}}} \times 100 = \frac{1.3 \times 10^{-4}}{0.50} \times 100 = 0.026\%$ is less than 5% ✓

a) Eq []'s: $[\text{H}_3\text{O}^+] = [\text{ClO}^-] = \mathbf{1.3 \times 10^{-4} \text{ M}}$; $[\text{HClO}] = 0.50 - 1.3 \times 10^{-4} = \mathbf{0.50 \text{ M}}$

b) $\text{pH} = -\log [\text{H}_3\text{O}^+]$ $\text{pH} = -\log 1.3 \times 10^{-4}$ **pH = 3.89**

2. a) Find the K_a of a 1.25 M solution of nitrous acid, $\text{HNO}_2(\text{aq})$. The pH of this solution is measured to be 1.62. **HNO₂ is WA.**



	$\text{HNO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NO}_2^-$			
I	1.25		0	0
C	-x		+x	+x
E	1.25 - x		x	x

$\text{pH} = 1.62$ $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$ $[\text{H}_3\text{O}^+] = 10^{-1.62} = 2.4 \times 10^{-2} \text{ M}$; thus, $x = 2.4 \times 10^{-2}$

Eq []'s: $[\text{HNO}_2] = 1.25 - 2.4 \times 10^{-2} = \mathbf{1.23 \text{ M}}$; $[\text{H}_3\text{O}^+] = [\text{NO}_2^-] = \mathbf{2.4 \times 10^{-2} \text{ M}}$

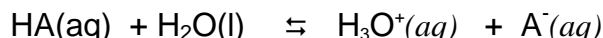
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} \quad K_a = \frac{x^2}{1.25 - x} = \frac{(2.4 \times 10^{-2})^2}{1.23} = \mathbf{4.7 \times 10^{-4}}$$

- Note: don't ignore -x since we know x ($x = 10^{-\text{pH}}$); we only assume x is small to avoid solving a quadratic (or polynomial) formula!**

- b) What is the %dissociation for this $\text{HNO}_2(\text{aq})$ solution?

$$\% \text{dissociation} = \frac{[\text{H}_3\text{O}^+]}{[\text{HA}]_{\text{initial}}} \times 100 \quad \% \text{dissociation} = \frac{2.4 \times 10^{-2}}{1.25} \times 100 = \mathbf{1.9\%}$$

3. A 0.200 M solution of a weak acid is 9.4% dissociated. Use this information to calculate $[\text{H}_3\text{O}^+]$, $[\text{A}^-]$, $[\text{HA}]$ and K_a .



	HA(aq) + H ₂ O(l)	⇌	H ₃ O ⁺ (aq) + A ⁻ (aq)
I	0.200		0
C	-x		+x
E	0.200 - x		x

$$\% \text{ionization} = \frac{x}{[\text{HA}]_{\text{in}}} \times 100 \quad 9.4\% = \frac{x}{0.200} \times 100$$

$$x = \frac{0.200(9.4)}{100} = 0.0188 \quad x = 0.019$$

$$\text{Eq []'s: } [\text{HA}] = 0.200 - 0.019 = \mathbf{0.181 \text{ M}}; [\text{H}_3\text{O}^+] = [\text{A}^-] = \mathbf{1.9 \times 10^{-2} \text{ M}}$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{x^2}{0.200 - x} = \frac{(0.019)^2}{0.181} = \mathbf{2.0 \times 10^{-3}}$$

4. Calculate the pH of a 0.50 M dimethylamine (CH₃)₂NH solution. K_b = 5.4 × 10⁻⁴
(CH₃)₂NH is WB

WB hydrolysis: (CH₃)₂NH₂ + H₂O ⇌ (CH₃)₂NH₃⁺ + OH⁻

	(CH ₃) ₂ NH ₂ + H ₂ O	⇌	(CH ₃) ₂ NH ₃ ⁺ + OH ⁻
I	0.50		0
C	-x		+x
E	0.50 - x		x

$$K_b = \frac{[(\text{CH}_3)_2\text{NH}_3^+][\text{OH}^-]}{[(\text{CH}_3)_2\text{NH}_2]} \quad \text{assume } x \text{ is small! } 5.4 \times 10^{-4} = \frac{x^2}{0.50}$$

$$x = \sqrt{(5.4 \times 10^{-4})(0.50)} = \sqrt{2.7 \times 10^{-4}} = 1.64 \times 10^{-2} \quad x = [\text{OH}^-] = 1.64 \times 10^{-2} \text{ M}$$

$$\text{check assumption: } \frac{x}{[\text{B}]_{\text{initial}}} \times 100 = \frac{1.64 \times 10^{-2}}{0.50} \times 100 = 3.28\% \text{ is less than } 5\% \checkmark$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log 1.64 \times 10^{-2} = 1.79 \quad \text{pH} = 14 - \text{pOH} = 14 - 1.79 = \mathbf{12.21}$$

5. A 0.065 M solution of methylamine, CH₃NH₂, has a pH of 11.70. What is K_b for CH₃NH₂?

WB hydrolysis: CH₃NH₂ + H₂O ⇌ CH₃NH₃⁺ + OH⁻ pH = 11.70

	CH ₃ NH ₂ + H ₂ O	⇌	CH ₃ NH ₃ ⁺ + OH ⁻
I	0.065		0
C	-x		+x
E	0.065 - x		x

$$\text{pOH} = 14 - \text{pH} \quad \text{pOH} = 14 - 11.70 = 2.30$$

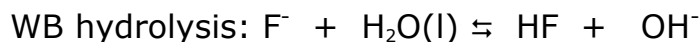
$$[\text{OH}^-] = 10^{-2.30} = 5.0 \times 10^{-3} \text{ M}; \text{ thus } x = 5.0 \times 10^{-3} \text{ M}$$

$$\text{Eq []'s: } [\text{CH}_3\text{NH}_2] = 0.065 - 0.0050 = \mathbf{0.060 \text{ M}}; [\text{OH}^-] = [\text{CH}_3\text{NH}_3^+] = \mathbf{5.0 \times 10^{-3} \text{ M}}$$

$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} \quad K_b = \frac{(5.0 \times 10^{-3})^2}{0.060} = \mathbf{4.2 \times 10^{-4}}$$

- Note we don't ignore -x since we find x from the pOH; we only assume x is small to avoid solving a quadratic (or polynomial) formula!

6. Is KF an acidic, basic or neutral salt? Write the hydrolysis reaction and calculate the pH of a 0.10 M KF solution. $K_a(\text{HF}) = 3.5 \times 10^{-4}$



	$\text{F}^- + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HF} + \text{OH}^-$			
I	0.10		0	0
C	-x		+x	+x
E	$0.10 - x$		x	x

1:1 ratio so $[\text{F}^-] = [\text{KF}] = 0.10 \text{ M}$ K_a for HF = 3.5×10^{-4}

$$K_a \times K_b = 1 \times 10^{-14} \quad K_b \text{ for F}^- = \frac{1 \times 10^{-14}}{3.5 \times 10^{-4}} = 2.86 \times 10^{-11}$$

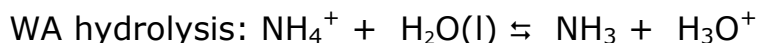
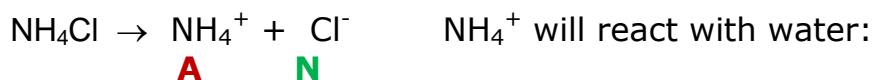
$$K_b = \frac{[\text{HF}][\text{OH}^-]}{[\text{F}^-]} \quad \text{Assume } x \text{ is small! } 2.86 \times 10^{-11} = \frac{x^2}{0.10}$$

$$x^2 = (2.86 \times 10^{-11})(0.10) \quad x = \sqrt{(2.86 \times 10^{-11})(0.10)} = 1.69 \times 10^{-6} \quad x = [\text{OH}^-] = \underline{1.69 \times 10^{-6} \text{ M}}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log [1.69 \times 10^{-6}] = 5.77 \quad \text{pH} = 14 - 5.77 = \underline{8.23 \text{ basic}} \checkmark$$

***Note: You must use K_b instead of K_a because F^- acts as a weak base in water; for WB problems remember that $x = [\text{OH}^-]$, so $-\log x$ gives us pOH not pH!**

7. Is NH_4Cl an acidic, basic, or neutral salt? Write the hydrolysis reaction and calculate the pH of a 0.10 M NH_4Cl solution. $K_b(\text{NH}_3) = 1.8 \times 10^{-5}$



	$\text{NH}_4^+ + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$			
I	0.10		0	0
C	-x		+x	+x
E	$0.10 - x$		x	x

1:1 ratio so $[\text{NH}_4^+] = [\text{NH}_4\text{Cl}] = 0.10 \text{ M}$ K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$

$$K_a = \frac{1 \times 10^{-14}}{K_b} \quad K_a \text{ for } \text{NH}_4^+ = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$$

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} \quad \text{Assume } x \text{ is small! } 5.56 \times 10^{-10} = \frac{x^2}{0.10}$$

$$x^2 = (5.56 \times 10^{-10})(0.10) \quad x = \sqrt{5.56 \times 10^{-11}} = 7.46 \times 10^{-6} \quad x = [\text{H}_3\text{O}^+] = \underline{7.46 \times 10^{-6} \text{ M}}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \quad \text{pH} = -\log \underline{7.46 \times 10^{-6}} \quad \underline{\text{pH} = 5.13 \text{ acidic}} \checkmark$$

***Note: You must use K_a instead of K_b because NH_4^+ acts as a weak acid in water; for WA problems remember that $x = [\text{H}_3\text{O}^+]$, so $-\log x$ gives us pH!**