Exercise 14.6
Ionization of Bases

Name: ______________________

Date:________________ Per:________

DIRECTIONS: Complete the following in the space provided.

1. Write the dissociation reaction and the corresponding K_b equilibrium expression for each of the following substances acting as bases in water.

   a. PO_4^{3-} $PO_4^{3-}(aq) + HOH(l) \rightleftharpoons OH^-(aq) + HPO_4^{2-}(aq)$ $K_b = [OH^-][HPO_4^{2-}] / [PO_4^{3-}]$

   b. HPO_4^{2-} $HPO_4^{2-}(aq) + HOH(l) \rightleftharpoons OH^-(aq) + H_2PO_4^-(aq)$ $K_b = [OH^-][H_2PO_4^-] / [HPO_4^{2-}]$

   c. NH_3 $NH_3(aq) + HOH(l) \rightleftharpoons OH^-(aq) + NH_4^+(aq)$ $K_b = [OH^-][NH_4^+] / [NH_3]$

   d. CN^- $CN^-(aq) + HOH(l) \rightleftharpoons OH^-(aq) + HCN(aq)$ $K_b = [OH^-][HCN] / [CN^-]$

2. Order the following bases from strongest to weakest. NO_3^-, H_2O, NH_3, CH_3NH_2 : CH_3NH_2, NH_3, H_2O, NO_3^-

3. Which is the stronger base?

   a. NO_3^- or NH_3 NH_3

   b. H_2O or NH_3 NH_3

   c. OH^-, or NH_3 OH^-

   d. NH_3 or CH_3NH_2 CH_3NH_2

4. Thallium(I) hydroxide is a **strong base** used in the synthesis of some organic compounds. Calculate the pH of a solution containing **2.48 g TlOH per liter**.

   $\frac{2.48\text{ g TlOH}}{} \bigg| \frac{1\text{ mol TlOH}}{221.387\text{ g TlOH}} = 0.0112\text{ mol TlOH}$

   Since TlOH is a strong base, it dissociates completely – the concentration of OH- will equal the concentration of the compound from which is dissociates.

   0.0112 mol / 1 L = 0.0112 *M* OH^- pOH = -log[OH^-] = -log(0.0112) = 1.951 pH = 14.000 – pOH = 12.049

5. Calculate [OH^-], pOH and pH for each of the following:

   a. 0.25 M NaOH
   NaOH is a strong base so [OH^-] = [NaOH]
   [OH^-] = 0.25 *M*
   pOH = -log[OH^-] = 0.60
   pH = 14.00 - pOH = 13.40

   b. A solution containing 25 g KOH per liter.
   KOH is a strong base so [OH^-] = [KOH]
   $\frac{25\text{ g KOH}}{} \bigg| \frac{1\text{ mol KOH}}{56.105\text{ g KOH}} = 0.45\text{ mol KOH}$
   [OH^-] = 0.45 *M*
   pOH = -log[OH^-] = 0.35
   pH = 14.00 - pOH = 13.65

   c. 0.00040 M $Ba(OH)_2$
   $Ba(OH)_2$ is a strong base. [OH^-] = 2[$Ba(OH)_2$]
   [OH^-] = 0.00080 *M*
   pOH = -log[OH^-] = 3.10
   pH = 14.00 - pOH = 10.90

   d. A solution containing 150.0 g NaOH per liter.
   NaOH is a strong base so [OH^-] = [NaOH]
   $\frac{150.0\text{ g NaOH}}{} \bigg| \frac{1\text{ mol NaOH}}{39.997\text{ g NaOH}} = 3.750\text{ mol NaOH}$
   [OH^-] = 3.750 *M*
   pOH = -log[OH^-] = -0.5740
   pH = 14.00 - pOH = 14.5740

6. What major species are present in 0.150 M solutions of each of the following bases? Calculate the [OH^-] and the pH of each of these solutions.

   a. KOH
   Major Species: $K^+(aq)$, $OH^-(aq)$
   KOH is a strong base so [OH^-] = [KOH]
   [OH^-] = 0.150 *M*
   pOH = -log[OH^-] = 0.824
   pH = 14.00 - pOH = 13.176

   b. CsOH
   Major Species: $Cs^+(aq)$, $OH^-(aq)$
   CsOH is a strong base so [OH^-] = [CsOH]
   [OH^-] = 0.150 *M*
   pOH = -log[OH^-] = 0.824
   pH = 14.00 - pOH = 13.176

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c. NH_3
Major Species: NH_3(aq), NH_4^+(aq), OH^-(aq)
Weak Base, so need K_b to find pH

d. Pyridine (C_5H_5N)
Major Species: C_5H_5N(aq), $C_5H_5NH^+$(aq), OH^-(aq)
Weak Base, so need K_b to find pH

7. What are the major species present in the following mixtures of bases?

a. 0.050 M NaOH and 0.050 M KOH ______ Na^+(aq), K^+(aq), OH^-(aq)

b. 0.050 M NaOH and 0.50 M NH_3 ______ Na^+(aq), OH^-(aq), NH_3(aq), NH_4^+(aq)

c. 0.0010 M $Ba(OH)_2$ and 0.020 M NaOH ______ Ba^+(aq), Na^+(aq), OH^-(aq)

8. For the reaction of hydrazine (N_2H_4) in water, $H_2NNH_2(aq) + H_2O(l) \rightleftharpoons H_2NNH_3^+(aq) + OH^-(aq)$, K_b is 3.0×10^{-6}. Calculate the concentrations of all species and the pH of a 2.0 M solution of hydrazine in water.

R	H_2NNH_2(aq) +	H_2O(l) ⇌	$H_2NNH_3^+$(aq) +	OH^-(aq)
I	2.0	N/A	0	0
C	-x	N/A	+x	+x
E	2.0-x	N/A	x	x

K_b is small enough to disregard the change in concentration of H_2NNH_2

$$K_b = \frac{[H_2NNH_3^+][OH^-]}{[H_2NNH_2]} \rightarrow 3.0 \times 10^{-6} = \frac{x^2}{2.0} \rightarrow x = 0.0024 = [OH^-]$$

pOH = -log[OH-] = -log[0.00240] = 2.620

pH = 14.000 – pOH = **11.380**

9. Calculate the $[OH^-]$, $[H^+]$, and pH of 0.20 M solutions of each of the following amines.

a. Ethylamine ($K_b = 5.6 \times 10^{-4}$)

R	B(aq) +	H_2O(l) ⇌	BH^+(aq) +	OH^-(aq)
I	0.20	N/A	0	0
C	-x	N/A	+x	+x
E	0.20-x	N/A	x	x

K_b is small enough to disregard the change in concentration of the initial base.

$$K_b = \frac{[BH^+][OH^-]}{[B]} \rightarrow 5.6 \times 10^{-4} = \frac{x^2}{0.20} \rightarrow x = 0.011 = [OH^-]$$

pOH = -log[OH-] = =log[0.011] = 1.96

pH = 14.000 – pOH = **12.04**

b. Diethylamine ($K_b = 1.3 \times 10^{-3}$)

R	B(aq) +	H_2O(l) ⇌	BH^+(aq) +	OH^-(aq)
I	0.20	N/A	0	0
C	-x	N/A	+x	+x
E	0.20-x	N/A	x	x

K_b is small enough to disregard the change in concentration of the initial base.

$$K_b = \frac{[BH^+][OH^-]}{[B]} \rightarrow 1.3 \times 10^{-3} = \frac{x^2}{0.20} \rightarrow x = 0.016 = [OH^-]$$

pOH = -log[OH-] = =log[0.016] = 1.80

pH = 14.000 – pOH = **12.20**

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c. Triethylamine ($K_b = 4.0 \times 10^{-4}$)

R	B(aq)	+	$H_2O(l)$	⇌	$BH^+(aq)$	+	$OH^-(aq)$
I	0.20		N/A		0		0
C	-x		N/A		+x		+x
E	0.20-x		N/A		x		x

K_b is small enough to disregard the change in concentration of the initial base.

$$K_b = \frac{[BH^+][OH^-]}{[B]} \rightarrow 4.0 \times 10^{-4} = \frac{x^2}{0.20} \rightarrow x = 0.0089 = [OH^-]$$

pOH = -log[OH-] = -log[0.0089] = 2.05

pH = 14.000 – pOH = **11.95**

(These weren't on your worksheet, but they're good practice.)

10. Calculate the percent ionization in each of the following solutions.

a. 0.010 M NH_3

R	$NH_3(aq)$	+	$H_2O(l)$	⇌	$NH_4^+(aq)$	+	$OH^-(aq)$
I	0.010		N/A		0		0
C	-x		N/A		+x		+x
E	0.010-x		N/A		x		x

K_a of $NH_4^+ = 5.6 \times 10^{-10}$, so K_b of $NH_3 = 1.8 \times 10^{-5}$

K_b is small enough to disregard the change in concentration of $[NH_3]$.

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \rightarrow 1.8 \times 10^{-5} = \frac{x^2}{0.010} \rightarrow x = 4.\underline{2}4 \times 10^{-4} = [OH^-] \rightarrow \frac{4.\underline{2}4 \times 10^{-4}}{0.010} \times 100 = \boxed{4.2\%}$$

b. 0.10 M hydroxylamine ($K_a = 9.333 \times 10^{-7}$)

R	$NH_2OH(aq)$	+	$H_2O(l)$	⇌	$NH_3OH^+(aq)$	+	$OH^-(aq)$
I	0.10		N/A		0		0
C	-x		N/A		+x		+x
E	0.10-x		N/A		x		x

K_a of $NH_3OH^+ = 9.333 \times 10^{-7}$, so K_b of $NH_2OH = 1.072 \times 10^{-8}$

K_b is small enough to disregard the change in concentration of $[NH_2OH]$.

$$K_b = \frac{[NH_3OH^+][OH^-]}{[NH_2OH]} \rightarrow 1.072 \times 10^{-8} = \frac{x^2}{0.10} \rightarrow x = 3.\underline{2}7 \times 10^{-5} = [OH^-] \rightarrow \frac{3.\underline{2}7 \times 10^{-5}}{0.10} \times 100 = \boxed{0.033\%}$$

c. 0.10 M methylamine

R	$CH_3NH_2(aq)$	+	$H_2O(l)$	⇌	$CH_3NH_3^+(aq)$	+	$OH^-(aq)$
I	0.10		N/A		0		0
C	-x		N/A		+x		+x
E	0.10-x		N/A		x		x

K_a of $CH_3NH_3^+ = 2.29 \times 10^{-11}$, so K_b of $CH_3NH_2 = 4.37 \times 10^{-4}$

K_b is small enough to disregard the change in concentration of $[CH_3NH_2]$.

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} \rightarrow 4.37 \times 10^{-4} = \frac{x^2}{0.10} \rightarrow x = 0.00661 = [OH^-] \rightarrow \frac{0.00661}{0.10} \times 100 = \boxed{6.61\%}$$