

Name: Key  
 Blk: \_\_\_\_\_ Date: \_\_\_\_\_

CHEMISTRY 12  
 ACID BASES UNIT

Lesson #13

Calculations Involving  $K_b$  and pOH

Recall that Strong BASES IONIZE 100 %, therefore the concentration of the strong base will equal the concentration of  $\text{OH}^-$  in solution!!

HOWEVER, Weak bases DO NOT IONIZE 100% in water, therefore we must use an ICE TABLE to determine the  $[\text{OH}^-]$  that is actually present in solution!

Generic Equation for a WEAK BASE in water:



There are THREE types of problems that you can solve associated with a weak base:

**Type 1.** Given the concentration of the weak base, solve for the pH (or pOH)

**Example 1.** What is the pH ( and pOH) for a 0.10 M solution of  $\text{NH}_3$ ?

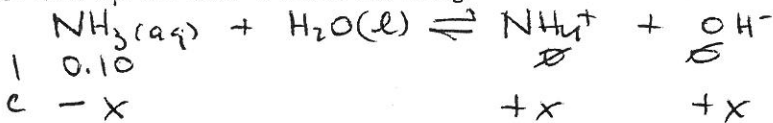
Step 1 . Write out the ionization equation with water



Step 2. Write out the  $K_b$  expression, identify the  $K_b$  value

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \rightarrow K_b(\text{NH}_3) = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.8 \times 10^{-5}$$

Step 3. Set up an ICE TABLE for  $\text{NH}_3$



Step 4. Plug values into  $K_b$  and solve for x

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.10]} = \sqrt{1.8 \times 10^{-5} \cdot 0.10} = \sqrt{x^2}$$

$$1.3 \times 10^{-3} = x$$

Step 5. Use x to solve for pH

$$[\text{OH}^-] = 1.3 \times 10^{-3} \text{ M} \rightarrow \text{pOH} = -\log(1.3 \times 10^{-3}) = 2.87$$

$$\text{pH} = 14.000 - 2.87 = \boxed{11.13}$$

**Type 2.** Given the pH (or pOH) and the initial concentration of the weak base, solve for the  $K_b$  value.

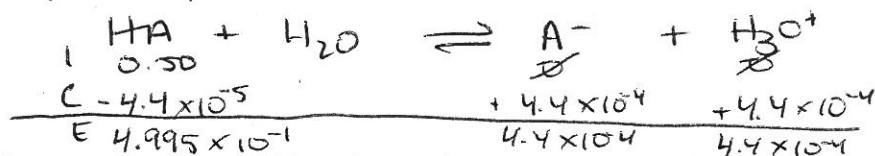
let  $0.10 - x \cong 0.10$

**Example 2.** If the pOH of a 0.50 M solution of an unknown weak acid is 10.64, determine the  $K_b$  for  $A^-$

Step 1. Write out the generic ionization with water of the acid



Step 2. Set up an ICE TABLE for the weak acid



Step 3. Use the pOH to determine pH and then fill in the values of the ICE TABLE

$$pOH = 10.64 \quad \therefore \quad pH = 14.000 - 10.64 = 3.36$$

$$\therefore [H_3O^+]_E = \text{antilog}(-3.36) = 4.4 \times 10^{-4}$$

Step 4. Use the equilibrium values to calculate the  $K_a$

$$K_a = \frac{(4.4 \times 10^{-4})^2}{(4.995 \times 10^{-1})} = 3.8 \times 10^{-7}$$

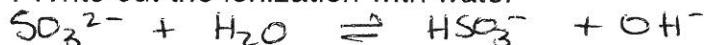
Step 5. Use the  $K_w = K_a (\text{conjugate acid}) \times K_b (\text{conjugate base})$  to solve for  $K_b$

$$\therefore K_b = \frac{1.00 \times 10^{-14}}{3.8 \times 10^{-7}} = \boxed{2.6 \times 10^{-8}}$$

**Type 3.** Given pH (or pOH), determine the initial concentration of the weak base.

**Example 3.** What concentration of  $SO_3^{2-}$  is required to produce a pH of 9.69?

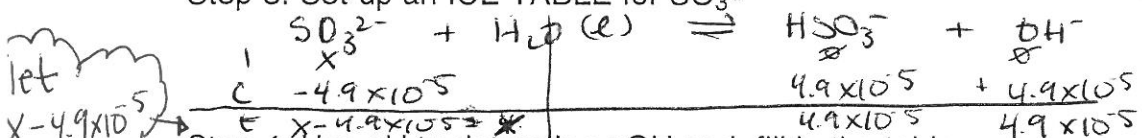
Step 1. Write out the ionization with water



Step 2. Write out the  $K_b$  expression, calculate  $K_b$  value

$$K_b = \frac{[HSO_3^-][OH^-]}{[SO_3^{2-}]} = K_b(SO_3^{2-}) = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7}$$

Step 3. Set up an ICE TABLE for  $SO_3^{2-}$



let  $x - 4.9 \times 10^{-5} \approx x$

Step 4. Use pH to determine pOH and fill in the table

$$pH = 9.69 \quad \therefore \quad pOH = 14.000 - 9.69 = 4.31 \quad [OH^-]_E = \text{antilog}(-4.31) = 4.9 \times 10^{-5}$$

Step 5. Use  $K_b$  to solve for initial concentration

$$1.0 \times 10^{-7} = \frac{(4.9 \times 10^{-5})^2}{x}$$

$$x = \frac{(4.9 \times 10^{-5})^2}{1.0 \times 10^{-7}}$$

$$x = 2.4 \times 10^{-2} \text{ M}$$

$$\boxed{[SO_3^{2-}] = 2.4 \times 10^{-2} \text{ M}}$$

**Seatwork/Homework:** Exercises 84-87

**PLO's:** Part of M3 and M5 for  $K_b$