Name: Key
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CHEMISTRY 12
ACID BASES UNIT
Lesson #13

Calculations Involving Kb and pOH

Recall that Strong BASES IONIZE 100 %, therefore the concentration of the strong base will equal the concentration of OH- in solution!!

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

Generic Equation for a WEAK BASE in water:

$$A^- + H_2O \longleftrightarrow HA + OH^-$$

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 NH_3?
Step 1 . Write out the ionization equation with water

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

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

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \longrightarrow K_b(NH_3) = \frac{1.00\times10^{-14}}{5.6\times10^{-10}} = 1.8\times10^{-5}$$

Step 3. Set up an ICE TABLE for NH_3

	$NH_{3(aq)}$ +	$H_2O(l)$ $\rightleftharpoons$	NH_4^+ +	OH^-
I	0.10		$\emptyset$	$\emptyset$
C	$-x$		$+x$	$+x$
E	$0.10 - x \cong 0.10$		x	x

let $0.10 - x \cong 0.10$

Step 4. Plug values into Kb and solve for x

$$1.8\times10^{-5} = \frac{[x][x]}{[0.10]} = \sqrt{1.8\times10^{-5}\cdot 0.10} = \sqrt{x^2}$$

$$1.3\times10^{-3} = x$$

Step 5. Use x to solve for pH

$$[OH^-] = 1.3\times10^{-3}\ M \longrightarrow pOH = -\log(1.3\times10^{-3}) = 2.87$$

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

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

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

$$HA + H_2O \rightleftharpoons A^- + H_3O^+$$

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

	HA	+ H_2O	$\rightleftharpoons$ A^-	+ H_3O^+
I	0.50		0	0
C	-4.4×10^{-5}		$+4.4\times10^{-4}$	$+4.4\times10^{-4}$
E	4.995×10^{-1}		4.4×10^{-4}	4.4×10^{-4}

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

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

$$\therefore [H_3O^+]_E = \text{antilog}(-3.36) = 4.4\times10^{-4}$$

Step4. Use the equilibrium values to calculate the K_a

$$Ka = \frac{(4.4\times10^{-4})^2}{(4.995\times10^{-1})} = 3.8\times10^{-7}$$

Step 5. Use the **Kw = Ka (conjugate acid) x Kb(conjugate base)** to solve for K_b

$$\therefore Kb = \frac{1.00\times10^{-14}}{3.8\times10^{-7}} = \boxed{2.6\times10^{-8}}$$

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

$$SO_3^{2-} + H_2O \rightleftharpoons HSO_3^- + OH^-$$

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

$$Kb = \frac{[HSO_3^-][OH^-]}{[SO_3^{2-}]} \qquad Kb(SO_3^{2-}) = \frac{1.00\times10^{-14}}{1.0\times10^{-7}} = 1.0\times10^{-7}$$

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

	SO_3^{2-}	+ $H_2O(l)$	$\rightleftharpoons$ HSO_3^-	+ OH^-
I	x		0	0
C	-4.9×10^{-5}		4.9×10^{-5}	$+4.9\times10^{-5}$
E	$x-4.9\times10^{-5}\approx x$		4.9×10^{-5}	4.9×10^{-5}

let $x-4.9\times10^{-5} \cong x$

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

$$pH = 9.69 \quad \therefore pOH = 14.000 - 9.69 = 4.31 \qquad [OH^-]_E = \text{antilog}(-4.31) = 4.9\times10^{-5}$$

Step 5. Use K_b to solve for initial concentration

$$1.0\times10^{-7} = \frac{(4.9\times10^{-5})^2}{x}$$

$$x = \frac{(4.9\times10^{-5})^2}{1.0\times10^{-7}}$$

$$x = 2.4\times10^{-2}\ M$$

$$\boxed{[SO_3^{2-}]_I = 2.4\times10^{-2}\ M}$$

Seatwork/Homework: Exercises 84-87

PLO's: Part of M3 and M5 for K_b