What is the pH of 0.10M $Ca(OH)_2$? (strong base)

$Ca(OH)_2 + \cancel{H_2O(l)} \rightarrow Ca^{2+} + 2OH^-$

0.10M   0.10   0.20M

$pOH = -\log[OH^-]$

$= -\log(0.20)$

$pOH = 0.70$

$pH = pKw - pOH$

$= 14.000 - 0.70$

$\boxed{pH = 13.30}$

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 RICE table to determine the [OH-] that is actually present in solution!

Generic Equation for a WEAK BASE in water:

$$A^- + H_2O(l) \leftrightarrow HA + OH^-$$

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

$NH_3 + H_2O(l) \rightleftharpoons NH_4^+ + OH^-$

Step 2. Write out the K_b expression, identify the K_b value $\rightarrow K_b = \frac{K_w}{K_a\ (NH_4^+)}$

$K_b = \boxed{\frac{[NH_4^+][OH^-]}{[NH_3]}} = \frac{1.00 \cdot 10^{-14}}{5.6 \cdot 10^{-10}} = \boxed{1.8 \cdot 10^{-5}}$

Step 3.

R	NH_3	+ $\cancel{H_2O(l)}$	$\rightleftharpoons NH_4^+$	+ OH^-
I	0.10		Ø	Ø
C	−x		+x	+x
E	0.10		x	x

let $0.10 - x \approx$

Step 4. Plug values into Kb and solve for x

$0.10\,(1.8 \cdot 10^{-5}) = \left(\frac{x^2}{\cancel{0.10}}\right) \cancel{0.10} \rightarrow \sqrt{(1.8 \cdot 10^{-5})(0.10)} = \sqrt{x^2}$

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

$1.3 \cdot 10^{-3}\ M = [OH^-]$

Step 5. Use x to solve for pH

$pOH = -\log(1.3 \cdot 10^{-3})$

$\boxed{pOH = 2.87}$

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

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

$HA + H_2O(l) \rightleftharpoons A^- + H_3O^+$; $Ka = \frac{[A^-][H_3O^+]}{[HA]}$

Step 2

R	HA	$+ \cancel{H_2O(l)}$	$\rightleftharpoons A^-$	$+ H_3O^+$
I	0.50		Ø	Ø
C	$-4.4\cdot10^{-4}$		$+4.4\cdot10^{-4}$	$+4.4\cdot10^{-4}$
E	0.50		$4.4\cdot10^{-4}$	$4.4\cdot10^{-4}$

0.50 − 0.00044 = 0.50

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

$[H_3O^+]_E = ?$

pH = pKw − pOH
= 14.000 − 10.64
pH = 3.36

$[H_3O^+]_E$ = antilog (−pH)
= antilog (−3.36)
$4.4\cdot10^{-4}$ M H_3O^+

Step4. Use the equilibrium values to calculate the K_a

$$Ka = \frac{(4.4\cdot10^{-4})^2}{0.50} = 3.8\cdot10^{-7}$$

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

$$\frac{1.00\cdot10^{-14}}{3.8\cdot10^{-7}} = \boxed{2.6\cdot10^{-8} \quad Kb \ A^-}$$

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

$SO_3^{2-} + H_2O(l) \longrightarrow HSO_3^- + OH^-$

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

$$\boxed{Kb = \frac{[HSO_3^-][OH^-]}{[SO_3^{2-}]} = 1.0\cdot10^{-7}}$$

$$Kb = \frac{Kw}{Ka(HSO_3^-)} = \frac{1.00\cdot10^{-14}}{1.0\cdot10^{-7}} = 1.0\cdot10^{-7}$$

R	SO_3^{2-}	$+ \cancel{H_2O(l)}$	$\rightleftharpoons HSO_3^-$	$+ OH^-$
I	X		Ø	Ø
C	$-4.9\cdot10^{-5}$		$+4.9\cdot10^{-5}$	$+4.9\cdot10^{-5}$
E	$x - 4.9\cdot10^{-5} \approx X$		$4.9\cdot10^{-5}$	$4.9\cdot10^{-5}$

$$x\left(\frac{(4.9\cdot10^{-5})^2}{x}\right) = (1.0\cdot10^{-7})x$$

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

pOH = 14.000 − 9.69
= 4.31

$[OH^-]_E$ = antilog (− 4.31)
= $4.9\cdot10^{-5}$

$$\frac{(4.9\cdot10^{-5})^2}{1.0\cdot10^{-7}} = x$$

Step 5. Use K_b to solve for initial concentration

$2.4\cdot10^{-2} = x$

$$\boxed{[SO_3^{2-}]_I = 2.4\cdot10^{-2} M}$$

Seatwork/Homework: Exercises 84-87, 89 + 91

PLO's: Part of M3 and M5 for K_b