

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



$$\text{pOH} = -\log[\text{OH}^-]$$

$$= -\log(0.20)$$

$$\text{pOH} = 0.70$$

CHEMISTRY 12
ACID BASES UNIT
Lesson #13

$$\text{pH} = \text{pK}_w - \text{pOH}$$

$$= 14.000 - 0.70$$

$$\text{pH} = 13.30$$

Calculations Involving K_b 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 $[\text{OH}^-]$ that is actually present in solution!

Generic Equation for a WEAK BASE in water:

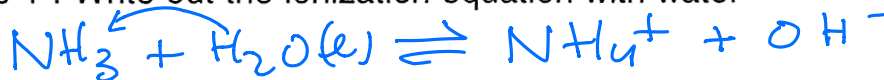


There are THREE types of problems that you can solve associated with a weak base:

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 NH_3 ?

Step 1 . Write out the ionization equation with water



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

$$K_b = \frac{K_w}{K_a(\text{NH}_4^+)}$$

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

Step 3. RICE table:

R	NH_3	H_2O	\rightleftharpoons	NH_4^+	OH^-
I	0.10			0	0
C	-x			+x	+x
E	0.10			x	x

let $0.10 - x \approx 0.10$

Step 4. Plug values into K_b and solve for x

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

Step 5. Use x to solve for pH

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

$$\text{pOH} = 2.87$$

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

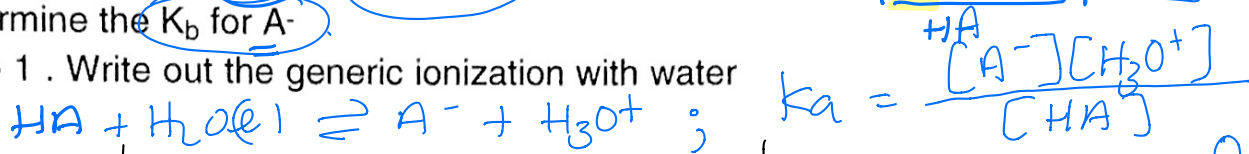
$$1.3 \cdot 10^{-3} \text{ M} = [\text{OH}^-]$$

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

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



Step 2

R	HA	+ H ₂ O	⇌	A ⁻	+ H ₃ O ⁺
I	0.50			0	0
C	-4.4 · 10 ⁻⁴			+4.4 · 10 ⁻⁴	+4.4 · 10 ⁻⁴
E	0.50			4.4 · 10 ⁻⁴	4.4 · 10 ⁻⁴

$$\begin{array}{r} 0.50 \\ - 0.00044 \\ \hline 0.50 \end{array}$$

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

[H₃O⁺]_E = ? pH = pK_w - pOH = 14.000 - 10.64 = 3.36

[H₃O⁺]_E = antilog(-pH) = antilog(-3.36) = 4.4 · 10⁻⁴ M H₃O⁺

Step 4. Use the equilibrium values to calculate the K_a

$$K_a = \frac{(4.4 \cdot 10^{-4})^2}{0.50} = 3.8 \cdot 10^{-7}$$

Step 5. Use the K_w = K_a (conjugate acid) x K_b (conjugate base) to solve for K_b

$$\frac{1.00 \cdot 10^{-14}}{3.8 \cdot 10^{-7}} = 2.6 \cdot 10^{-8} \quad K_b \quad A^-$$

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

Example 3. What concentration of SO₃²⁻ 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

$$K_b = \frac{K_w}{K_a(HSO_3^-)} = \frac{1.00 \cdot 10^{-14}}{1.0 \cdot 10^{-7}} = 1.0 \cdot 10^{-7}$$

$$\frac{[HSO_3^-][OH^-]}{[SO_3^{2-}]} = 1.0 \cdot 10^{-7}$$

R	SO ₃ ²⁻	+ H ₂ O	⇌	HSO ₃ ⁻	+ OH ⁻
I	X			0	0
C	-4.9 · 10 ⁻⁵			+4.9 · 10 ⁻⁵	+4.9 · 10 ⁻⁵
E	4.9 · 10 ⁻⁵			4.9 · 10 ⁻⁵	4.9 · 10 ⁻⁵

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

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 · 10⁻⁵

Step 5. Use K_b to solve for initial concentration

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

$$2.4 \cdot 10^{-2} = x$$

$$[SO_3^{2-}]_I = 2.4 \cdot 10^{-2} M$$

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

PLO's: Part of M3 and M5 for K_b