# *Acid base equilibria 2*

*The ionization constant for $NH_4OH$ is 1.8 x $10^{-5}$. Calculate the concentration of $OH^-$ ions in a 1.0 molar solution of $NH_4OH$.*

## Solution:

The ionization constant ($K_b$) is defined as the concentration of $OH^-$ ions times the concentration of the conjugate acid ions of a given base divided by the concentration of an unionized base. For a base (BA),

$$K_b := \frac{(B^-) \cdot (A^+)}{(BA)}$$

Where $K_b$ is the ionized constant, ($B^-$) is the concentration of ionized base ions, ($A^+$) is the concentration of the conjugate acid, and (BA) is the concentration of the unionized base. The $K_b$ for $NH_4OH$ is stated as

$$K_b = \frac{(NH_4^+) \cdot (OH^-)}{NH_4OH} = 1.8 \times 10^{-5}$$

When $NH_4OH$ is ionized, one $NH_4^+$ ion is formed and one $OH^-$ ion is formed,

$$NH_4OH \rightleftharpoons NH_4^+ + OH^-$$

The concentrations of each ions are equal.

$$(NH_4^+) = (OH^-)$$

The concentration of the unionized base is decreased when ionization occurs. The new concentration is equal to the concentration of $OH^-$ subtracted from the concentration of $NH_4OH$.

$$(NH_4OH) = 1.0 - (OH^-)$$

Since ($OH^-$) is small relative to 1.0, one may assume that 1.0 - ($OH^-$) is approximately equal to 1.0.

$$(NH_4OH) = 1.0 - (OH^-) \approx 1.0$$

Using the assumption and the fact that $(ON^-) = (NH_4^+)$, $K_b$ can be rewritten as

$$K_b = \frac{(OH^-) \cdot (OH^-)}{1.0} = 1.8 \times 10^{-5}$$

Solution for $(OH^-)$:

$$\frac{(OH^-) \cdot (OH^-)}{1.0} = 1.8 \times 10^{-5}$$

$$(OH^-)^2 = 1.8 \times 10^{-5}$$

$$(OH^-) = \sqrt[2]{1.8 \times 10^{-5}} = 4.2 \times 10^{-3}$$