TYPE 1 (weak BASE)

(only a weak BASE put in solution)

For a weak base (B) in water:

$$B + H_2O \rightleftharpoons OH^- + BH^+$$

$K_b = \dfrac{[OH^-][BH^+]}{[B]}$ and let C_B be the analytical concentration (label) of the weak base

Our Equilibrium "Chart"

	[B]	**[OH⁻]**	**[BH⁺]**
initial	C_B	0	0
change	$-x$	$+x$	$+x$
final	$C_B - x$	x	x

$$K_b = \frac{x^2}{(C_B - x)} \qquad \textbf{Equation 1.1}$$

$$K_b C_B - K_b(x) = x^2$$

$$0 = x^2 + K_b(x) - K_b C_B$$

$$\boxed{[OH^-] = [BH^+] = x = \frac{-K_b + \sqrt{K_b^2 + 4K_b C_B}}{2}}$$

This equation is the <u>exact</u> solution for calculating the [OH-] for any solution made by mixing a concentration of ONLY a weak base in water. (What I have designated a "Type 1" problem)

If we are using a REASONABLE concentration for B (somewhere between 1.0 M and 0.05 M) AND K_b is small enough (say less than 10^{-4}) we can make the following assumption:

$C_B - x \cong C_B$ all this is saying is that x is so small compared to C_B, it doesn't change it

So that Equation 1.1 shown above becomes

$$K_b = \frac{x^2}{C_B}$$

and solving:

$$\boxed{[OH^-] = [BH^+] = x = \sqrt{K_b C_B} \qquad \textbf{Equation 1.2}}$$

This equation is the <u>approximate</u> solution for calculating the $[OH^-]$ for any weak base with $K_b < 10^{-4}$ when put in water.