pH Calculation of a Weak Base Solution

What is the pH of a 1.0M NH_3 solution?

step 1: | H_2O NH_3 NH_4^+ OH^{1-} | NH_3, NH_4^+, OH^{1-}, H_2O

step 2

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^{1-} \quad K_{b_{NH_3}} = 1.8\times10^{-5}$$

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^{1-} \quad K_w = 1.0\times10^{-14}$$

step 3

Since $K_{b_{NH_3}} \gg K_w$, assume all OH^{1-} are from NH_3

step 4

$$K_b = \frac{[NH_4^+]_E^1 [OH^{1-}]_E^1}{[NH_3]_E^1}$$

$[NH_3]_I = 1.0M$
$[NH_4^+]_I = [OH^{1-}]_I = 0M$

step 5

	NH_3	OH^{1-}	NH_4^+
I	1.0M	0M	0M
C	$-x$	$+x$	$+x$
E	$1.0M - x$	x	x

Assume 5% rule
$1.0M - x \approx 1.0M$

$$\frac{1.341\times10^{-2}M}{1.0m}\cdot 100\% < 5\%$$

✓ check

$$1.8\times10^{-4}M = \frac{(x)(x)}{1.0M - x}$$

$$1.8\times10^{-4}M = \frac{x^2}{1.0M}$$

$$[NH_4^+]_E = [OH^{1-}]_E = x = \sqrt{(1.0M)(1.8\times10^{-4}M)} = 1.341\times10^{-2}M$$

$$[OH^{1-}]_E[H^+]_E = 1.0\times10^{-14}M^2 \Rightarrow [H^+]_E = \frac{1.0\times10^{-14}M^2}{[OH^{1-}]_E} = \frac{1.0\times10^{-14}M^2}{1.341\times10^{-2}M} = 7.457\times10^{-13}M$$

$$pH = -\log[H^+] = -\log(7.457\times10^{-13}M) = -(-12.127) = 12.127$$

$$\boxed{pH = 12.13}$$