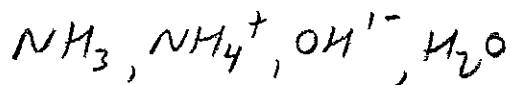
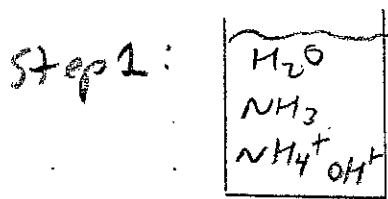
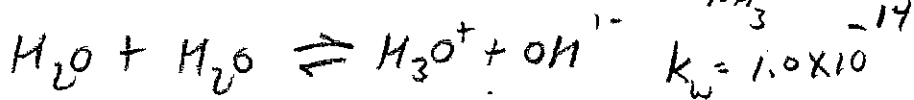
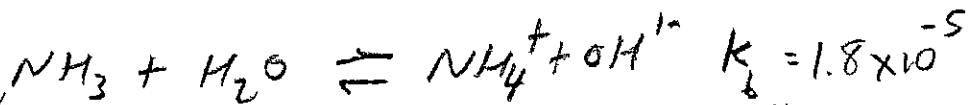


pH Calculation of a Weak Base Solution

What is the pH of a 1.0M NH_3 solution?



Step 2



Step 3

Since $K_{b, \text{NH}_3} \gg K_w$, assume all OH^- are from NH_3

Step 4

$$K_b = \frac{[\text{NH}_4^+]_E [\text{OH}^-]_E}{[\text{NH}_3]_E} \quad \begin{matrix} [\text{NH}_3]_E = 1.0\text{M} \\ [\text{NH}_4^+]_E = [\text{OH}^-]_E = x \end{matrix}$$

Step 5

	NH_3	OH^-	NH_4^+
I	1.0M	0M	0M
C	-x	+x	+x
E	1.0M-x	x	x

Assume 5% rule

$1.0\text{M} - x \approx 1.0\text{M}$

$\frac{1.341 \times 10^{-2}\text{M}}{1.0\text{M}} \cdot 100\% < 5\%$

✓ check

$$1.8 \times 10^{-4} \text{M} = \frac{(x)(x)}{1.0\text{M} - x}$$

$$1.8 \times 10^{-4} \text{M} = \frac{x^2}{1.0\text{M}}$$

$$[\text{NH}_4^+]_E = [\text{OH}^-]_E = x = \sqrt{(1.0\text{M})(1.8 \times 10^{-4}\text{M})} = 1.341 \times 10^{-2}\text{M}$$

$$[\text{OH}^-]_E [\text{H}^+]_E = 1.0 \times 10^{-14} \text{M}^2 \Rightarrow [\text{H}^+]_E = \frac{1.0 \times 10^{-14} \text{M}^2}{[\text{OH}^-]_E} = \frac{1.0 \times 10^{-14} \text{M}^2}{1.341 \times 10^{-2}\text{M}} = 7.457 \times 10^{-13}\text{M}$$

$$\text{pH} = -\log [\text{H}^+] = -\log (7.457 \times 10^{-13}\text{M}) = -(-12.127) = 12.127$$

$\text{pH} = 12.13$