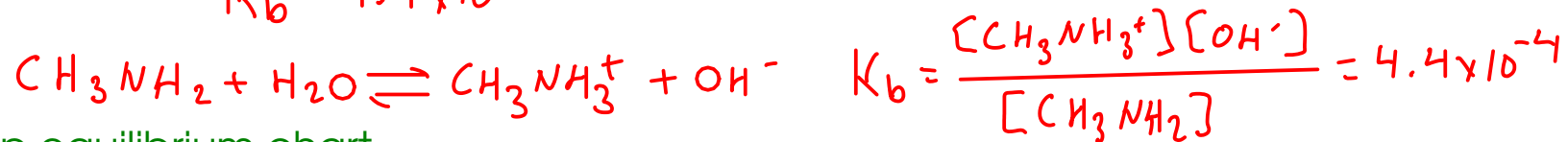


Find the pH of 0.17 M methylamine.

On page A-14, we find a BASE ionization constant for methylamine. It's therefore a WEAK BASE ... $K_b = 4.4 \times 10^{-4}$



Set up equilibrium chart ...

Species	[Initial]	Δ	[Equilibrium]
CH_3NH_3^+	0	+x	x
OH^-	0	+x	x
CH_3NH_2	0.17	-x	0.17-x

Let "x" equal the change in methylammonium ion concentration ...

$$\frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = \frac{(x)(x)}{(0.17-x)} = 4.4 \times 10^{-4}$$

$$\frac{x^2}{0.17-x} = 4.4 \times 10^{-4}$$

$x \ll 0.17$, so $0.17-x \approx 0.17$

$$\frac{x^2}{0.17} = 4.4 \times 10^{-4}$$

$$x = 0.0086486993 = [\text{OH}^-]$$

$$\text{pOH} = -\log_{10} [\text{OH}^-] = 2.06$$

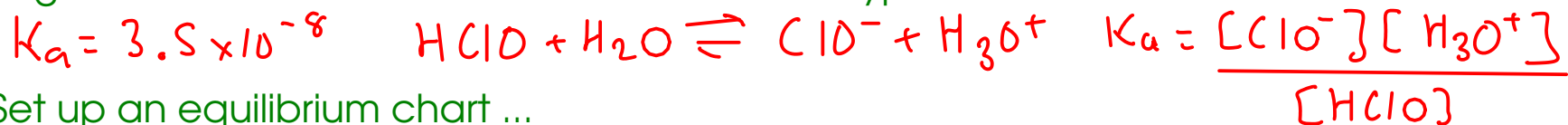
Since $\text{pH} + \text{pOH} = 14.00$

$$\text{pH} + 2.06 = 14.00$$

$$\text{pH} = 11.94$$

2 Find the pH of 0.11 M hypochlorous acid

Page A-13 lists an acid ionization constant for hypochlorous acid. It's a WEAK ACID.



Set up an equilibrium chart ...

Species	[Initial]	Δ	[Equilibrium]
ClO^-	0	+x	x
H_3O^+	0	+x	x
HClO	0.11	-x	0.11 - x

Let "x" equal the change in hypochlorite ion concentration ...

$$\frac{[\text{ClO}^-][\text{H}_3\text{O}^+]}{[\text{HClO}]} = \frac{(x)(x)}{0.11 - x} = 3.5 \times 10^{-8}$$

$$\frac{x^2}{0.11 - x} = 3.5 \times 10^{-8}$$

$x \ll 0.11, \text{ so}$
 $0.11 - x \approx 0.11$

$$\frac{x^2}{0.11} = 3.5 \times 10^{-8}$$

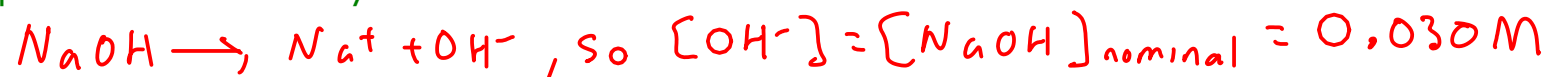
$$x = 6.204836823 \times 10^{-5} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+] = 4.21$$

³ Find the pH of 0.030 M sodium hydroxide.

Sodium hydroxide is a Group IA hydroxide ... a common STRONG BASE

Since the STRONG base COMPLETELY ionizes, the hydroxide concentration will simply equal the sodium hydroxide concentration...



$$[\text{OH}^-] = 0.030 \text{ M}$$

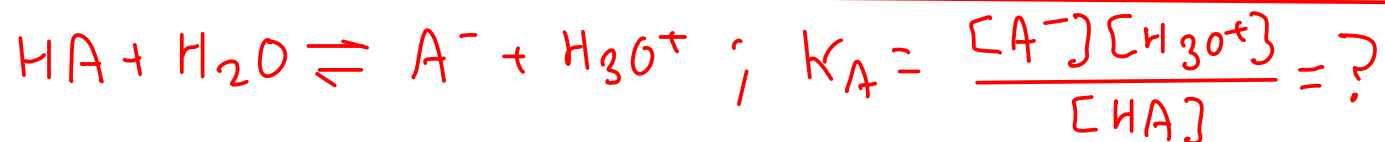
$$\text{pOH} = -\log_{10}(0.030) = 1.52$$

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} + 1.52 = 14.00$$

$$\boxed{\text{pH} = 12.48}$$

4 An 0.15 M solution of monoprotic acid has a pH of 2.80 at 25 C. Find the K_a of the acid



Write an equilibrium chart to reduce the number of variables ...

Species	[Initial]	Δ	[Equilibrium]
A^-	0	+x	x
H_3O^+	0	+x	x
HA	0.15	-x	0.15 - x

Let "x" equal the change in A^- concentration...

$$\frac{[A^-][H_3O^+]}{[HA]} = \frac{x^2}{0.15 - x} = K_a$$

We still have two variables. "x", and K_a . To solve the problem, we must get "x" another way.

We can use the pH to find hydronium ion concentration, which is equal to "x" (see the equilibrium chart above!)

$$pH = 2.80$$
$$[H_3O^+] = 10^{-pH}$$

$$[H_3O^+] = 10^{-2.80} = 0.0615848932 = x$$

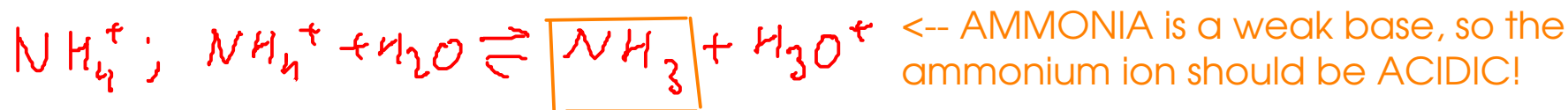
Plug value of "x" into the equilibrium expression ...

$$\frac{(0.0615848932)^2}{0.15 - 0.0615848932} = 1.7 \times 10^{-5} = K_a$$

5

Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate (FW=80.052 g/mol) solid into enough water to make 250. mL of solution.

What's the nature of ammonium nitrate? $\text{NH}_4\text{NO}_3 \rightarrow \text{NH}_4^+ + \text{NO}_3^-$



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

Ka value for ammonium ion? Page A-13 doesn't have it, but page A-14 has Kb for the conjugate, ammonia:

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5}$$

$$\text{So } K_{a, \text{NH}_4^+} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$$

To solve the equilibrium problem, we need to find the nominal concentration of our ammonium nitrate.

$$80.052 \text{ g NH}_4\text{NO}_3 = \text{mol NH}_4\text{NO}_3$$

$$3.00 \text{ g NH}_4\text{NO}_3 \times \frac{\text{mol NH}_4\text{NO}_3}{80.052 \text{ g NH}_4\text{NO}_3} = 0.0374756408 \text{ mol NH}_4\text{NO}_3$$

$$[\text{NH}_4\text{NO}_3]_{\text{nominal}} = \frac{0.0374756408 \text{ mol NH}_4\text{NO}_3}{0.250 \text{ L}} =$$

$$= 0.1499025633 \text{ M NH}_4\text{NO}_3$$



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.56 \times 10^{-10}$$

Species	[Initial]	Δ	[Equilibrium]
NH_3	0	+X	X
H_3O^+	0	+X	X
NH_4^+	0.14990	-X	0.14990 - X

Solve ...

$$\frac{x^2}{0.14990 - x} = 5.56 \times 10^{-10}$$

$$\downarrow x \ll 0.14990$$

$$\frac{x^2}{0.14990} = 5.56 \times 10^{-10}$$

$$x = 9.13 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = 5.04$$

... seems reasonable for a weakly acidic salt at moderate concentration.