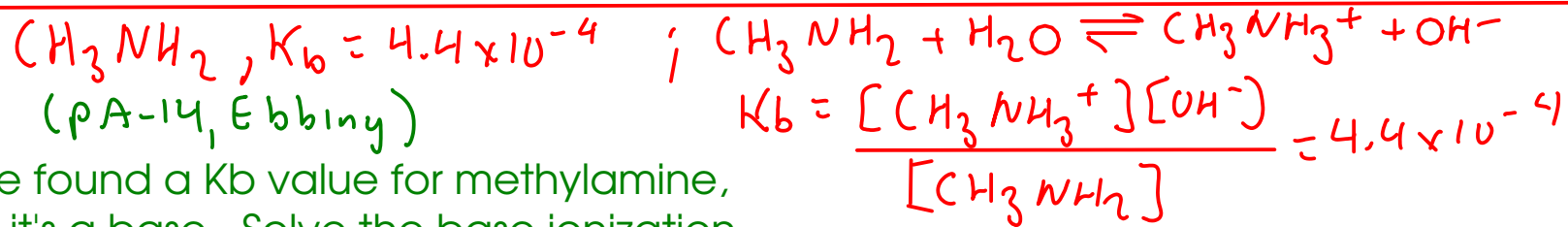


Find the pH of 0.17 M methylamine.



We found a K_b value for methylamine, so it's a base. Solve the base ionization equilibrium to find HYDROXIDE ion concentration.

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

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

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

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

Assume $x \ll 0.17$
 $0.17 - x \approx 0.17$

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

$$x = 0.0086486993 \text{ M OH}^-$$

$$\text{pOH} = 2.06$$

$$\text{pH} = 14.00 - 2.06 = \boxed{11.94}$$

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

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

2 Find the pH of 0.11 M hypochlorous acid



(pA-13, Ebbing)

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

We have an acid, so we need to solve the acid ionization equilibrium of HClO.

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

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

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

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

$x \ll 0.11$
 \downarrow
 $0.11 - x \approx 0.11$

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

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

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

³ Find the pH of 0.030 M sodium hydroxide.

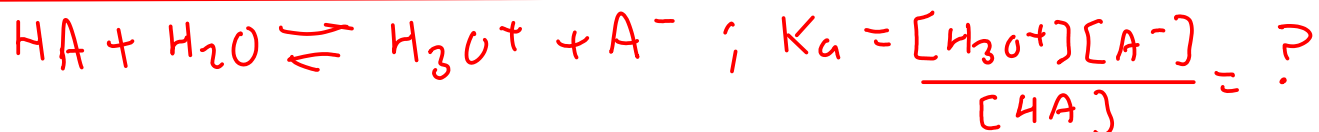
Strong base! It completely ionizes: $\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$
 $[\text{OH}^-] = [\text{NaOH}]_{\text{initial}} = 0,030 \text{ M OH}^-$

The concentration of hydroxide ion in solution is set by the base ... in this case it just equals whatever the concentration of the NaOH is.

$$\text{pOH} = -\log_{10}(0,030) = 1,52$$

$$\text{pH} = 14,00 - 1,52 = \boxed{12,48} \quad (\text{pH} + \text{pOH} = 14,00)$$

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



In this case, we are ASKED to find K_a ... and therefore we can't look it up in our book. Set up an equilibrium chart to reduce the number of variables in the K_a expression.

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

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

$$\frac{(x)(x)}{(0.15-x)} = K_a$$

$$\frac{x^2}{0.15-x} = K_a$$

We have two variables, "x" and " K_a ". If we can find out what "x" is, we can use it to get K_a . What is "x"?

We know that the solution has a pH of 2.80, and that gives us the hydronium ion concentration ... which is "x"

$$x = [H_3O^+], [H_3O^+] = 10^{-2.80} = 0.0015848932 \text{ M}$$

$([H_3O^+] = 10^{-\text{pH}})$

Plug into expression for K_a ...

$$\frac{(0.0015848932)^2}{0.15 - 0.0015848932} = K_a = 1.7 \times 10^{-5}$$

Find the pH of a solution made of 25 mL of 0.10 M hydrofluoric acid mixed with 75 mL of 0.075 M NaF.

This is a BUFFER solution. It's made of a weak acid (HF) and its conjugate base (fluoride ion). We can save time by using the HENDERSON-HASSELBALCH EQUATION!

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{basic species}]}{[\text{acidic species}]} \right)$$

$$K_{a, \text{HF}} = 6.8 \times 10^{-4}, \text{p}K_a = 3.17$$

(PA-13, Ebbing)

Since we mixed these solutions to make the buffer, find the new concentrations of HF and fluoride ion.

$$\text{HF}: M_1 V_1 = M_2 V_2 \quad (0.10 \text{ M})(25 \text{ mL}) = M_2 (100. \text{ mL})$$

$$0.025 \text{ M HF} = M_2$$

$$\text{F}^-: M_1 V_1 = M_2 V_2 \quad (0.075 \text{ M})(75 \text{ mL}) = M_2 (100 \text{ mL})$$

$$0.05625 \text{ M F}^- = M_2$$

$$\text{pH} = 3.17 + \log \left(\frac{0.05625 \text{ M}}{0.025 \text{ M}} \right) = \boxed{3.52}$$