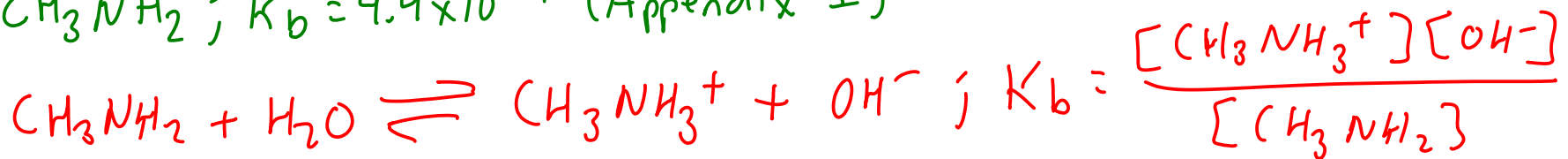
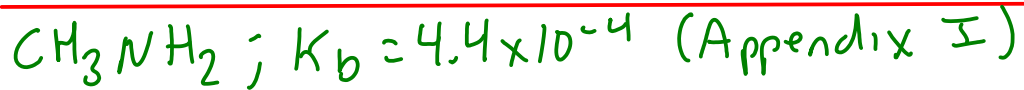


Find the pH of 0.17 M methylamine.



Solve for hydroxide 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 $x \ll 0.17$
↓

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

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

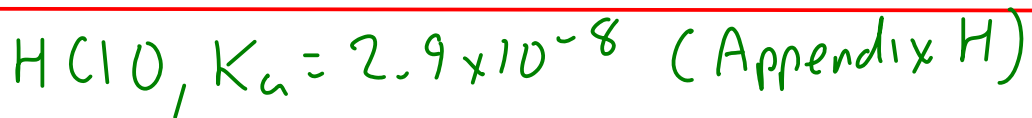
$$\text{pOH} = 2.06$$

$$\boxed{\text{pH} = 11.94}$$

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

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

2 Find the pH of 0.11 M hypochlorous acid



Solve equilibrium for hydronium ion concentration.

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

$$\frac{(x)(x)}{0.11 - x} = 2.9 \times 10^{-8}$$

Assume $x \ll 0.11$

$$\frac{x^2}{0.11} = 2.9 \times 10^{-8}$$

$$x = 5.648008499 \times 10^{-5} = [\text{H}_3\text{O}^+]$$
$$\text{pH} = \boxed{4.25}$$

³ Find the pH of 0.030 M sodium hydroxide.

Strong base. NaOH completely ionizes in water:



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

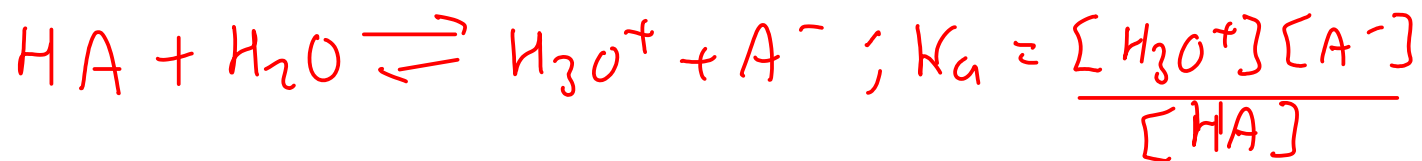
$$\text{pOH} = 1.52$$

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

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

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

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



Set up a chart to reduce the number of variables.

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.

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

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

Use the pH to find the value of "x".

$$[H_3O^+] = x = 10^{-2.80} \quad ([H_3O^+] = 10^{-pH})$$

$$x = 0.0015848932$$

Plug "x" back into K_a equation.

$$\frac{(0.0015848932)^2}{0.15 - 0.0015848932} = K_a$$

$$1.7 \times 10^{-5} = K_a$$

What is the pH of a solution that contains both 0.15 M formic acid and 0.10 M sodium formate?

This is a buffer, since it contains significant concentrations of a weak acid (formic acid) and its conjugate base (formate ion). Use the Henderson-Hasselbalch equation.

$$\text{pH} = \text{pK}_a + \log \left(\frac{[\text{basic species}]}{[\text{acidic species}]} \right) \leftarrow \begin{array}{l} 0.10 \text{ M} \\ 0.15 \text{ M} \end{array}$$

$$\begin{array}{l} \uparrow \\ K_{a, \text{HCOOH}} = 1.8 \times 10^{-4} \text{ (Appendix H)} \\ \text{pK}_a = 3.74 \end{array}$$

$$\text{pH} = 3.74 + \log \left(\frac{0.10 \text{ M}}{0.15 \text{ M}} \right)$$

$$\text{pH} = \boxed{3.57}$$