Find the pH of 0.17 M methylamine.

CH_3NH_2 ; $K_b = 4.4 \times 10^{-4}$ (Appendix I)

$$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^- \; ; \; K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

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

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

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

$\downarrow$ Assume $x \ll 0.17$

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

$$x = 0.0086486993 = [OH^-]$$

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

$$\boxed{pH = 11.94} \qquad pH + pOH = 14.00$$

 Find the pH of 0.11 M hypochlorous acid

HClO, $K_a = 2.9 \times 10^{-8}$ (Appendix H)

$$HClO + H_2O \rightleftarrows H_3O^+ + ClO^- \qquad K_a = \frac{[H_3O^+][ClO^-]}{[HClO]}$$

Solve equilibriium 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 << 0.11

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

$x = 5.648008499 \times 10^{-5} = [H_3O^+]$

pH = $\boxed{4.25}$

3 Find the pH of 0.030 M sodium hydroxide.

Strong base. NaOH completely ionizes in water:

$NaOH \rightarrow Na^+ + OH^-$, so

$[OH^-] = 0.030\, M$

$pOH = 1.52$ $\qquad pOH = -\log_{10}[OH^-]$

$pH = \boxed{12.48}$ $\qquad pH + pOH = 14.00$

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

$$HA + H_2O \rightleftharpoons H_3O^+ + A^- \; ; \; K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

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".

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

$$x = 0.0015848932$$

Plug "x" back into Ka equation.

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

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

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.

$$pH = pK_a + \log\left(\frac{[\text{basic species}] \leftarrow 0.10\,M}{[\text{acidic species}] \leftarrow 0.15\,M}\right)$$

$K_{a,HCOOH} = 1.8\times10^{-4}$ (Appendix H)

$pK_a = 3.74$

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

$$pH = \boxed{3.57}$$