Find the pH of 0.17 M methylamine.

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

$$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^- \qquad K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} = 4.4 \times 10^{-4}$$

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 ...

$$\frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} = \boxed{\frac{(x)(x)}{(0.17-x)} = 4.4 \times 10^{-4}}$$

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

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

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

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

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

Since $pH + pOH = 14.00$

$pH + 2.06 = 14.00$

$$\boxed{pH = 11.94}$$

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.

$K_a = 3.5 \times 10^{-8}$ $\quad HClO + H_2O \rightleftharpoons ClO^- + H_3O^+ \quad K_a = \dfrac{[ClO^-][H_3O^+]}{[HClO]}$

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 ...

$$\frac{[ClO^-][H_3O^+]}{[HClO]} = \boxed{\frac{(x)(x)}{0.10 - x} = 3.5 \times 10^{-8}}$$

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

$x \ll 0.11$, so $0.11 - x \approx 0.11$

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

$$x = 6.204836823 \times 10^{-5} = [H_3O^+]$$

$$pH = -\log_{10}[H_3O^+] = \boxed{4.21}$$

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...

$NaOH \rightarrow Na^+ + OH^-$, so $[OH^-] = [NaOH]_{nominal} = 0.030\,M$

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

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

$pH + pOH = 14.00$

$pH + 1.52 = 14.00$

$\boxed{pH = 12.48}$

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

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

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...

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

We still have two variables. "x", and Ka. 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!)

$pH = 2.80$

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

$[H_3O^+] = 10^{-2.80} = 0.001584893\,2 = x$

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

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

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? $NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$

NO_3^-: $NO_3^- + H_2O \rightleftharpoons \boxed{HNO_3} + OH^-$ <-- NITRIC ACID is a strong acid, so the nitrate ion should be NEUTRAL

NH_4^+; $NH_4^+ + H_2O \rightleftharpoons \boxed{NH_3} + H_3O^+$ <-- AMMONIA is a weak base, so the ammonium ion should be ACIDIC!

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

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

$K_{b,NH_3} = 1.8 \times 10^{-5}$

So $K_{a,NH_4^+} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$

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

$80.052\ g\ NH_4NO_3 = mol\ NH_4NO_3$

$$3.00\ g\ NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.052\ g\ NH_4NO_3} = 0.0374756408\ mol\ NH_4NO_3$$

$$[NH_4NO_3]_{nominal} = \frac{0.0374756408\ mol\ NH_4NO_3}{0.250\ L} =$$

$= 0.149902563\ M\ NH_4NO_3$

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56 \times 10^{-10}$$

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

Solve ...

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

$$\downarrow x \ll 0.14990$$

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

$$x = 9.13 \times 10^{-6} = [H_3O^+]$$

$$\boxed{pH = 5.04}$$

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