1. For each of the following aqueous solutions, calculate the pH:
   a. Nitric acid (HNO₃), 0.25M
      For strong acids, assume essentially complete dissociation:
      \[ \text{HNO}_3 (aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+ (aq) + \text{NO}_3^- (aq) \]
      \[ [\text{H}_3\text{O}^+] = 0.25 \text{ M} ; \quad \text{pH} = -\log([\text{H}_3\text{O}^+]) = -\log(0.25 \text{ M}) = 0.6 \]
   b. 0.01M NaOH
      For strong bases, assume essentially complete dissociation:
      \[ \text{NaOH}(aq) \rightarrow \text{HO}^- (aq) + \text{Na}^+ (aq) \]
      \[ \text{pH} + \text{pOH} = 14; \quad \text{pOH} = -\log([\text{OH}^-]) = -\log(0.01 \text{ M}) = 2; \quad \text{pH} = 14 - 2 = 12. \]
   c. 0.1M HCl
      strong acid, \( \text{pH} = -\log (0.1 \text{ M}) = 1 \)
   d. 0.02 M ClICH₂CO₂H which is 10% dissociated
      weak acid:
      \[ \text{ClICH}_2\text{CO}_2\text{H} (aq) + \text{H}_2\text{O}(l) \Leftrightarrow \text{H}_3\text{O}^+ (aq) + \text{ClICH}_2\text{CO}_2^- (aq) \]
      Initial: \( 0.02 \text{ M} \) \( 1 \times 10^{-7} \text{ M} \) \( 0 \text{ M} \)
      Final: \( 0.02 - 0.002 \text{ M} \) \( 0.002 \text{ M} \) \( 0.002 \text{ M} \)
      \[ \text{pH} = -\log (0.002) = 2.69 \]
   e. 2M LiOH: strong base
      \[ \text{pH} + \text{pOH} = 14; \quad \text{pOH} = -\log([\text{OH}^-]) = -\log (2 \text{ M}) = -0.3; \quad \text{pH} = 14 - (-0.3) = 14.3 \]

2. Write ionization constant expressions for the following acids and bases:
   a. \( \text{NH}_4^+ \)  b. \( \text{CN}^- \)  c. \( \text{NO}_2^- \)  d. \( \text{HCOOH} \)

   a. \( \text{NH}_4^+ (aq) + \text{H}_2\text{O} (l) \Leftrightarrow \text{H}_3\text{O}^+ (aq) + \text{NH}_3(aq) \)
      \[ K_a = [\text{H}_3\text{O}^+]_{eq}[\text{NH}_3]_{eq} / [\text{NH}_4^+]_{eq} \]
   b. \( \text{CN}^- (aq) + \text{H}_2\text{O} (l) \Leftrightarrow \text{HCN} (aq) + \text{HO}^- (aq) \)
      \[ K_a = [\text{HCN}]_{eq}[\text{HO}^-]_{eq} / [\text{CN}^-]_{eq} \]
   c. \( \text{NO}_2^- (aq) + \text{H}_2\text{O} (l) \Leftrightarrow \text{HNO}_2 (aq) + \text{HO}^- (aq) \)
      \[ K_a = [\text{HNO}_2]_{eq}[\text{HO}^-]_{eq} / [\text{NO}_2^-]_{eq} \]
   d. \( \text{HCOOH} (aq) + \text{H}_2\text{O} (l) \Leftrightarrow \text{H}_3\text{O}^+ (aq) + \text{HCOO}^- (aq) \)
      \[ K_a = [\text{H}_3\text{O}^+]_{eq}[\text{HCOO}^-]_{eq} / [\text{HCOOH}]_{eq} \]

3. Calculate the [\( \text{H}_3\text{O}^+ \)] and [\( \text{OH}^- \)] and the pH at equilibrium that results when 0.01 mol of each of the following acids is added to sufficient water to produce a 1L solution (treat this problem as you would any typical equilibrium problem):
   (a) \( \text{HOCl} \) \( K_a = 3.7 \times 10^{-8} \)
      First, write out the dissolution equation:
      \[ \text{HOCl}(aq) + \text{H}_2\text{O}(l) \Leftrightarrow \text{H}_3\text{O}^+(aq) + \text{ClO}^-(aq) \]
      Initial \( 0.01 \text{ M} \) \( 1 \times 10^{-7} \text{ M} \) \( 0 \text{ M} \)
      Equilibrium: \( 0.01 \text{ M} - x \) \( 1 \times 10^{-7} \text{ M} + x \) \( x \)
      Since \( K_a = [\text{H}_3\text{O}^+]_{eq}[\text{ClO}^-]_{eq} / [\text{HOCl}]_{eq} = 3.7 \times 10^{-8} = (1 \times 10^{-7} + x)(x) / (0.01M-x) \) then we have:
3.7 x 10^{-8} = (1x10^{-7}x + x^2) / (0.01 \text{ mol-x}); \ x^2 + 1.37x10^7x - 3.7 x 10^{-10} = 0; \text{ use quadratic} \\
x = 1.92 \times 10^{-3} \text{ mol} \\
For 1 L, [\text{HOCl}] = 0.01 \text{ mol-x} = 0.0099 \ M \\
[\text{H}_3\text{O}^+] = 1x10^{-7} + x = 1.93 \times 10^{-5} \text{ M}; \text{ pH} = 4.71; \text{ pOH} = 9.28; \\
[\text{OH}] = 5.2 \times 10^{-10} \text{ M} \\
[\text{ClO}^-] = x = 1.92 \times 10^{-5} \text{ M} \\
(b) \text{ HCN} \quad (K_a = 7.2 \times 10^{-10}) \\
First, write out the dissolution equation: \\
HCN(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + CN^-(aq) \\
Initial: 0.01 \ M \quad 1x10^{-7} \ M \quad 0 \ M \\
Equilibrium: 0.01 \ M-x \quad 1x10^{-7} \ M +x \quad x \\
Since: K_a = [H_3O^+]_{eq}[CN^-]_{eq} / [HCN]_{eq} = 7.2 \times 10^{-10} = (1x10^{-7} + x)(x) / (0.01 \text{ M-x}) \\
Since K_a<<1x10^{-7}, \text{ we cannot assume } x \gg 1x10^{-7}, \text{ thus:} \\
7.2 \times 10^{-12} - 7.2x10^{-10}x = 1x10^{-7}x + x^2; \ x^2 + 1x10^{-7}x - 7.2 \times 10^{-12} = 0 \\
using quadratic, x = 2.63 \times 10^{-6} \ M \\
thus [\text{H}_3\text{O}^+] = 1x10^{-7} \text{ mol-x} = 2.73 \times 10^{-6} \ M; \text{ pH} = 5.56; [\text{OH}] = 3.6 \times 10^{-9} \ M \\
[\text{CN}^-] = x = 2.63 \times 10^{-6} \ M \\
[\text{HCN}] = 0.01 \text{ mol-x} = 0.0099 \ M \\
4. Calculate the [\text{H}_3\text{O}^+] and [\text{OH}] and the pH at equilibrium that results when 0.15 mol of each of the following bases is added to sufficient water to produce a 1L solution (treat this problem as you would any typical equilibrium problem): \\
(a) \text{NH}_3 \quad (K_b = 1.8 \times 10^{-5}) \\
\text{NH}_3(aq) + H_2O(l) \leftrightarrow \text{NH}_4^+(aq) + \text{HO}^-(aq) \\
Initial: 0.15 \ M \quad 0 \quad 1x10^{-7} \ M \\
Equilibrium: 0.15-x \quad x \quad 1x10^{-7} + x \\
K_b = [\text{NH}_4^+]_{eq}[\text{HO}^-]_{eq} / [\text{NH}_3]_{eq} \\
K_b = 1.8 \times 10^{-5} = (x)(1x10^{-7}+x) / (0.15 -x) \\
2.7x10^{-6} - 1.8x10^{-5}x = 1x10^{-7}x + x^2; \ x^2 + 1.81x10^{-5}x - 2.7 \times 10^{-6} = 0 \\
quadratic: x = 0.0016 \ M \\
[\text{NH}_4^+] = x = 0.0016 \ M \\
[\text{HO}^-] = 1x10^{-7} + x = 0.0016 \ M; \ [\text{H}_3\text{O}^+] = 6.25 \times 10^{-12} \ M; \text{ pH} = 11.2 \\
[\text{NH}_3] = 0.15-x = 0.1484 \ M \\
(b) \text{(CH}_3)_3\text{N} \quad (K_b = 6.3 \times 10^{-5}) \\
5. The dissociation constant for formic acid (HCOOH) at 25°C is 1.8 x 10^{-4} \\
(a) For a 0.1 M solution, calculate the percentage of the acid that is dissociated into ions. \\
Assume 1L of solution: \\
HCOOH (aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + HCOO^-(aq) \\
Initial: 0.1 \ M \quad 1x10^{-7} \ M \quad 0 \ M \\
Eq: 0.1 - x \quad 1x10^{-7} + x \quad x \\
K_a = 1.8 \times 10^{-4} = (x)(1x10^{-7} + x) / (0.1 - x) \\
Assume x \gg 1x10^{-7} \text{ mol}, then } (x)^2/(0.1 - x) = 1.8 \times 10^{-4}; \ x^2 = 1.8x10^{-4} - 1.8x10^{-4}x
x\(^2\) + 1.8 \times 10\(^{-4}\)x – 1.8 \times 10\(^{-5}\) = 0 \quad ; \quad \text{quadratic: } x = .00415

\text{percentage of acid dissociated} = \frac{x}{0.1} \times 100\% = 4.15\%

(b) Calculate the pH of a water sample containing 6.32 g of formic acid per liter

Molar mass of HCOOH = 46 g/mol; 6.32 g / 46 g/mol = 0.137 mol

\text{HCOOH (aq) + H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HCOO}^-(aq)

\begin{align*}
\text{Initial:} & \quad 0.137 \text{ M} \quad 1 \times 10^{-7} \text{ M} \quad 0 \text{ M} \\
\text{Eq.:} & \quad 0.137-x \quad x \quad x
\end{align*}

K\(_a\) = 1.8 \times 10^{-4} = \frac{x^2}{(0.137-x)}; \quad 2.466 \times 10^{-5} - 1.8 \times 10^{-4}x = x^2

x^2 + 1.8 \times 10^{-4}x – 2.46 \times 10^{-5} = 0. \quad \text{quadratic: } x = 0.0048

\text{[H}_3\text{O}^+] = x = 0.0048 \text{ M}; \quad \text{pH} = 2.3

6. Calculate the dissociation constant at 25°C for nitrous acid (HNO}_2\text{) in a 0.1M solution if it is known to be 6.3% dissociated.

6.3% of 0.1M = 0.1M x 0.063 = 0.0063M

\text{HNO}_2\text{ (aq) + H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{NO}_2^-\text{(aq)}

\begin{align*}
\text{Init:} & \quad 0.1 \text{ M} \quad 1 \times 10^{-7} \text{ M} \quad 0 \text{ M} \\
\text{Eq.:} & \quad 0.0937 \text{ M} \quad 0.0063 \text{ M} \quad .0063 \text{ M}
\end{align*}

K\(_a\) = \frac{[\text{H}_3\text{O}^+]_{\text{eq}} \cdot [\text{NO}_2^-]_{\text{eq}}}{[\text{HNO}_2]_{\text{eq}}} = \frac{(0.0063\text{ M})^2}{(0.0937\text{ M})} = 4.2 \times 10^{-4}\text{M}