pH of Polyprotic Acids

1. **Write the ionization reactions for carbonic acid. Given that $K_{a1} = 4.4 \times 10^{-7}$ and $K_{a2} = 4.7 \times 10^{-11}$, calculate the $[H_2CO_3]$, $[HCO_3^-]$, $[CO_3^-]$, and the pH of a 0.080 M solution of H_2CO_3.**

① $H_2CO_3 \rightleftharpoons HCO_3^- + H^+$ $K_{a_1} = 10^{-7}$

② $HCO_3^- \rightleftharpoons CO_3^{2-} + H^+$ $K_{a_2} = 10^{-11}$

	H_2CO_3	$\rightleftharpoons HCO_3^-$	$+ H^+$
	.080	0	0
	$-x$	$+x$	$+x$
	$.080 - x$	x	x

$4.4 \times 10^{-7} = \frac{x^2}{.080}$

$x = 1.9 \times 10^{-4}\ M$

$[H^+] = [HCO_3^-] = 1.9 \times 10^{-4}\ M$

$[H_2CO_3] = .080\ M$

	HCO_3^-	$\rightleftharpoons H^+$	$+ CO_3^{2-}$
	1.9×10^{-4}	1.9×10^{-4}	0
	$-x$	$+x$	$+x$
	$1.9 \times 10^{-4} - x$	$1.9 \times 10^{-4} + x$	x

$4.7 \times 10^{-11} = \frac{(1.9 \times 10^{-4})(x)}{1.9 \times 10^{-4}}$ → negligible

$x = 4.7 \times 10^{-11}$

$[CO_3^{2-}] = 4.7 \times 10^{-11}\ M$

$pH = -\log[(1.9 \times 10^{-4}) + (4.7 \times 10^{-11})]$ (4.7×10^{-11} negligible)

$= -\log 1.9 \times 10^{-4}$

$= 3.7$

2. **Write the ionization reactions for the first and second ionizations of sulfuric acid. Remember, sulfuric acid is a strong acid. Given that the K_a of HSO_4^- is 1.1×10^{-2}, calculate the $[H_3O^+]$, $[HSO_4^-]$, $[SO_4^{-2}]$and pH of a 0.021 M solution of sulfuric acid.**

① $H_2SO_4 \rightarrow H^+ + HSO_4^-$

	H_2SO_4	$\rightarrow H^+$	$+ HSO_4^-$
	.021	0	0
	$-.021$	$+.021$	$+.021$
	0	.021	.021

gone!

② $HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$

	HSO_4^-	$\rightleftharpoons H^+$	$+ SO_4^{2-}$
	.021	.021	0
	$-x$	$+x$	$+x$
	$.021 - x$	$.021 + x$	x

do not use 5% Rule!

$K_{a_2} = 1.1 \times 10^{-2} = \frac{(.021 + x)(x)}{.021 - x}$

$x = .0061\ M$

$[SO_4^{2-}] = .0061\ M$

$[HSO_4^-] = .021 - .0061 = .015\ M$

$[H^+] = .021 + .0061 = .027\ M$

$pH = -\log .027$

$= 1.57$

3. Given that $Ka_1 = 6.5 \times 10^{-2}$ and $Ka_2 = 6.1 \times 10^{-5}$, calculate the pH of a 1.40 M $H_2C_2O_4$ (oxalic acid) solution and the equilibrium concentrations of $H_2C_2O_4$, $HC_2O_4^-$, $C_2O_4^{-2}$, and OH^-.

① $H_2C_2O_4 \rightleftharpoons H^+ + HC_2O_4^-$

1.40	0	0
$-x$	$+x$	$+x$
$1.40-x$	x	x

$$6.5 \times 10^{-5} = \frac{(x)(x)}{1.40-x}$$

$x = .271\ M$

$[H^+] = [HC_2O_4^-] = .271\ M$

$[H_2C_2O_4] = 1.40 - .271 = 1.13\ M$

$HC_2O_4^- \rightleftharpoons H^+ + C_2O_4^{2-}$

.271	.271	0
$-x$	$+x$	$+x$
$.271-x$	$.271+x$	x

$$6.1 \times 10^{-5} = \frac{(.271)(x)}{.271}$$

$x = 6.1 \times 10^{-5}\ M$

$[C_2O_4^{2-}] = 6.1 \times 10^{-5}\ M$

$pH = -\log .271 = .567$

$[OH^-] = \frac{Kw}{[H^+]} = \frac{1 \times 10^{-14}}{.271} = 3.7 \times 10^{-14}\ M\ OH^-$

4. Calculate the $[H_2PO_4^-]$, $[HPO_4^{-2}]$, $[PO_4^{-3}]$, $[OH^-]$, and pH of a 6.0 M phosphoric acid solution. $Ka_1 = 7.5 \times 10^{-3}$, $Ka_2 = 6.2 \times 10^{-8}$, and $Ka_3 = 4.8 \times 10^{-13}$.

$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$

6.0	0	0
$-x$	$+x$	$+x$
$6.0-x$	x	x

$$7.5 \times 10^{-3} = \frac{x^2}{6.0-x}$$

$x = .208 = .21\ M = [H^+] = [H_2PO_4^-]$

$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2-}$

.21	.21	0
$-x$	$+x$	$+x$
$.21-x$	$.21+x$	x

$$6.2 \times 10^{-8} = \frac{(.21)(x)}{.21}$$

$x = 6.2 \times 10^{-8}\ M = [HPO_4^{2-}]$

$HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-}$

6.2×10^{-8}	.21	0
$-x$	$+x$	$+x$
$6.2 \times 10^{-8} - x$	$.21+x$	x

$$4.8 \times 10^{-13} = \frac{(.21)(x)}{6.2 \times 10^{-8}}$$

$x = 1.4 \times 10^{-19}\ M = [PO_4^{3-}]$

$pH = -\log .21 = .68$

$[OH^-] = 10^{-13.32} = 4.8 \times 10^{-14}\ M$