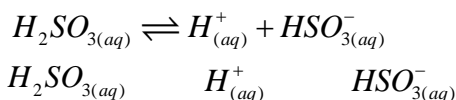# In Class Exercise for Chapter 17

1. Determine the pH for each of the following.

a.) 0.500 $\underline{M}$ H₂SO₃ with K_{a1} = 1.5 x 10⁻² and K_{a2} = 6.3 x 10⁻⁸

$$H_2SO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HSO^-_{3(aq)}$$

	$H_2SO_{3(aq)}$	$H^+_{(aq)}$	$HSO^-_{3(aq)}$
I	0.500	0	0
C	-x	+x	+x
E	0.500 - x	+x	+x

$$K_{a1} = \frac{\left[H^+_{(aq)}\right]\left[HSO^-_{(aq)}\right]}{\left[H_2SO_{3(aq)}\right]} = 1.5\times10^{-2} = \frac{x^2}{0.500-x} \rightarrow x^2 + 0.015x - 0.0075 = 0$$

$$\left[H^+_{(aq)}\right] = 7.94\times10^{-2}\underline{M} = \left[HSO^-_{(aq)}\right]$$

Since this is a polyprotic acid we are not done yet.

$$HSO^-_{3(aq)} \rightleftharpoons H^+_{(aq)} + SO^{2-}_{3(aq)}$$
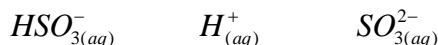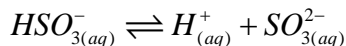
	$HSO^-_{3(aq)}$	$H^+_{(aq)}$	$SO^{2-}_{3(aq)}$
I	0.0794	0.0794	0
C	-x	+x	+x
E	0.0794 - x	0.0794 + x	+x

$$K_{a2} = \frac{\left[H^+_{(aq)}\right]\left[SO^{2-}_{3(aq)}\right]}{\left[HSO^-_{3(aq)}\right]} = 6.3\times10^{-8} = \frac{(0.0794+x)x}{0.0794-x} \sim x \text{ since } 10^{-8} \ll 10^{-2}$$

$$\left[H^+_{(aq)}\right] = 7.94\times10^{-2}\underline{M} \therefore pH = -\log(7.94\times10^{-2}) = 1.10$$

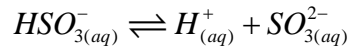Therefore, the second deprotonation does not significantly contribute to our pH.

b.) 0.300 $\underline{M}$ Na₂SO₃

$$SO^{2-}_{3(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HSO^-_{3(aq)}$$
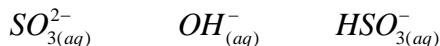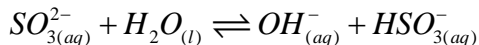
	$SO^{2-}_{3(aq)}$	$OH^-_{(aq)}$	$HSO^-_{3(aq)}$
I	0.300	0	0
C	-x	+x	+x
E	0.300 - x	+x	+x

$$K_b = \frac{K_w}{K_{a2}} = \frac{1.0\times10^{-14}}{6.3\times10^{-8}} = 1.59\times10^{-7}$$

$$K_b = \frac{\left[OH^-_{(aq)}\right]\left[HSO^-_{3(aq)}\right]}{\left[SO^{2-}_{3(aq)}\right]} = 1.59\times10^{-7} = \frac{x^2}{0.300-x} \sim \frac{x^2}{0.300} \text{ since } 10^{-7} \ll 10^{-1}$$

$$\left[OH^-_{(aq)}\right] = 2.18\times10^{-4}\underline{M} \therefore pH = 14 + \log(2.18\times10^{-4}\underline{M}) = 10.34$$

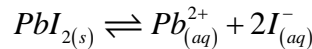c.) 0.300 L solution containing 0.500 M NaHSO$_3$ and 0.300 M Na$_2$SO$_3$

$$HSO_{3(aq)}^- \rightleftharpoons H_{(aq)}^+ + SO_{3(aq)}^{2-}$$

$$pH = pK_a + \log\left(\frac{\left[SO_{3(aq)}^{2-}\right]}{\left[HSO_{3(aq)}^-\right]}\right)$$

$$pH = -\log(6.3\times10^{-8}) + \log\left(\frac{0.300}{0.500}\right)$$

$$pH = 6.98$$

2. Determine the molar solubility for each of the following:

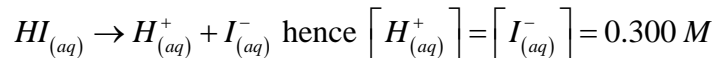a.) 0.500 $\underline{M}$ PbI$_2$ with K$_{sp}$ = 1.4 x 10^{-8}

$$PbI_{2(s)} \rightleftharpoons Pb_{(aq)}^{2+} + 2I_{(aq)}^-$$

	$Pb_{(aq)}^{2+}$	$I_{(aq)}^-$
I	0	0
C	+s	+2s
E	+s	+2s

$$K_{sp} = 1.4\times10^{-8} = s(2s)^2 = 4s^3$$

$$s = 0.0015\ M \quad \left[Pb_{(aq)}^{2+}\right] = 0.0015\ M \quad \& \quad \left[I_{(aq)}^-\right] = 0.0030\ M$$

b.) 0.300 $\underline{M}$ HI

$$HI_{(aq)} \rightarrow H_{(aq)}^+ + I_{(aq)}^- \text{ hence } \left[H_{(aq)}^+\right] = \left[I_{(aq)}^-\right] = 0.300\ M$$

c.) 1 L solution containing 0.500 $\underline{M}$ PbI$_2$ and 0.300 $\underline{M}$ HI

	$Pb_{(aq)}^{2+}$	$I_{(aq)}^-$
I	0	0.300
C	+s	+2s
E	+s	0.300+2s

$$K_{sp} = 1.4\times10^{-8} = s(0.300+2s)^2$$

since $10^{-8} \ll 10^{-1} \rightarrow 1.4\times10^{-8} \sim 0.0900s$ or $s = 1.56\times10^{-7}\ M$

ck assumption: $\dfrac{1.56\times10^{-7}}{0.300}\times100\% = 5\times10^{-5}\ \%$

$$\left[Pb_{(aq)}^{2+}\right] = 1.6\times10^{-7}\ M \quad \& \quad \left[I_{(aq)}^-\right] = 0.300\ M$$
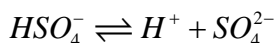
3. . Calculate the pOH and the concentrations of all species in a solution prepared by mixing equal volumes of 0.2 M HCl and 0.6 M H$_2$SO$_4$ (K$_{a2}$ = 1.2 x 10^{-2}).

we have two strong acids which completely dissociate:

$HCl \rightarrow H^+ + Cl^-$ and $H_2SO_4 \rightarrow H^+ + HSO_4^-$

Initially: $\left[H^+\right] = \dfrac{1\,L \times 0.2\,\frac{mol}{L} + 1\,L \times 0.6\,\frac{mol}{L}}{2L} = 0.4\,\frac{mol}{L}$ and $\left[HSO_4^-\right] = \dfrac{1\,L \times 0.6\,\frac{mol}{L}}{2L} = 0.3\,\frac{mol}{L}$

since is a polyprotic acid we use the equation below to determine eq:

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

	HSO_4^-	H^+	SO_4^{2-}
Initial	0.30	0.4	0
Change	-x	+x	+x
Eq	0.30 - x	0.4+x	+x

$K_a = \dfrac{\left[H^+\right]\left[SO_4^{2-}\right]}{\left[HSO_4^-\right]} = \dfrac{x(0.4+x)}{0.3-x} = 1.2 \times 10^{-2} \rightarrow x^2 + 0.412x - 0.0036 = 0$

$x = \dfrac{-0.412 \pm \sqrt{0.412^2 - 4(1)(-0.0036)}}{2} \rightarrow x = 0.0085M$ or $x = $ -0.420M

$\left[HSO_4^-\right] = 0.3 - x = 0.291M \rightarrow 0.3M$

$\left[H^+\right] = 0.4 + x = 0.409M \rightarrow 0.4M$

$\left[SO_4^{2-}\right] = 0.009M$

$pOH = 14 - pH = 14 + \log\left[H^+\right] = 13.60$

4. What is the pH at the endpoint for titration of 0.20 __M__ solutions of the following:

   a.) $C_5H_{11}N$ (piperidine – $K_b = 1.3$ x 10^{-3}) & HNO₃
   When we add equal amounts of acid and base they react completely to produce
   salt and water: $C_5H_{11}N + HNO_3 \rightarrow C_5H_{11}NH^+NO_3^-$ or $C_5H_{11}N + H^+ \rightarrow C_5H_{11}NH^+$
   Furthermore, addition of equal amounts will cut our concentrations in half. So,
   we start with the 0.10 __M__ $C_5H_{11}NH^+$ in water.

$C_5H_{11}NH^+ + H_2O \rightleftharpoons C_5H_{11}N + H_3O^+$
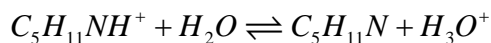
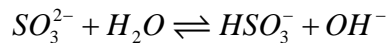|	$C_5H_{11}NH^+$	$H^+_{(aq)}$	$C_5H_{11}N$
I	0.10	0	0
C	-x	+x	+x
E	0.10 - x	+x	+x

$K_a = \dfrac{K_w}{K_b} = \dfrac{1.0 \times 10^{-14}}{1.3 \times 10^{-3}} = 7.7 \times 10^{-12}$

$K_a = 7.7 \times 10^{-12} = \dfrac{x^2}{0.10 - x} \sim \dfrac{x^2}{0.10}$ since $10^{-12} \ll 10^{-1}$

$\left[H^+\right] = 8.8 \times 10^{-7}\,\underline{M} \therefore pH = 6.06$

b.) $NaHSO_3$ ($K_a = 6.3 \times 10^{-8}$) & $NaOH$

 Similar to the last problem in which our acid HSO_3^- has been completely

 neutralized by the hydroxide. So we have 0.10 <u>M</u> SO_3^{2-} in water.

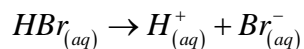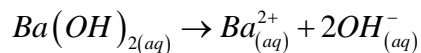$$SO_3^{2-} + H_2O \rightleftharpoons HSO_3^- + OH^-$$

	SO_3^{2-}	OH^-	HSO_3^-
I	0.10	0	0
C	-x	+x	+x
E	0.10 - x	+x	+x

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.3 \times 10^{-8}} = 1.6 \times 10^{-7}$$

$$K_b = 1.6 \times 10^{-7} = \frac{x^2}{0.10 - x} \sim \frac{x^2}{0.10} \text{ since } 10^{-7} << 10^{-1}$$

$$\left[OH^- \right] = 1.26 \times 10^{-4} \underline{M} \therefore pH = 14 + \log\left(1.26 \times 10^{-4}\right) = 10.10$$

c.) $Ba(OH)_2$ & HBr

$$Ba(OH)_{2(aq)} \rightarrow Ba^{2+}_{(aq)} + 2OH^-_{(aq)}$$

$$HBr_{(aq)} \rightarrow H^+_{(aq)} + Br^-_{(aq)}$$

the rxn is really btwn $H^+_{(aq)}$ and $OH^-_{(aq)}$

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} \quad pH = 7.00$$