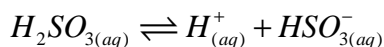


In Class Exercise for Chapter 15

1. Determine the pH for each of the following.

a.) 0.500 M H_2SO_3 with $K_{a1} = 1.5 \times 10^{-2}$ and $K_{a2} = 6.3 \times 10^{-8}$

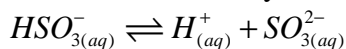


	$H_2SO_{3(aq)}$	$H_{(aq)}^+$	$HSO_{3(aq)}^-$
I	0.500	0	0
C	-x	+x	+x
E	0.500 - x	+x	+x

$$K_{a1} = \frac{[H_{(aq)}^+][HSO_{3(aq)}^-]}{[H_2SO_{3(aq)}]} = 1.5 \times 10^{-2} = \frac{x^2}{0.500 - x} \rightarrow x^2 + 0.015x - 0.0075 = 0$$

$$[H_{(aq)}^+] = 7.94 \times 10^{-2} \underline{M} = [HSO_{3(aq)}^-]$$

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



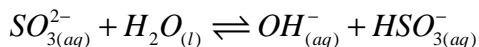
	$HSO_{3(aq)}^-$	$H_{(aq)}^+$	$SO_{3(aq)}^{2-}$
I	0.0794	0.0794	0
C	-x	+x	+x
E	0.0794 - x	0.0794 + x	+x

$$K_{a2} = \frac{[H_{(aq)}^+][SO_{3(aq)}^{2-}]}{[HSO_{3(aq)}^-]} = 6.3 \times 10^{-8} = \frac{(0.0794 + x)x}{0.0794 - x} \sim x \text{ since } 10^{-8} \ll 10^{-2}$$

$$[H_{(aq)}^+] = 7.94 \times 10^{-2} \underline{M} \therefore pH = -\log(7.94 \times 10^{-2}) = 1.10$$

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

b.) 0.300 M Na_2SO_3



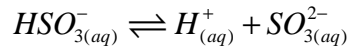
	$SO_{3(aq)}^{2-}$	$OH_{(aq)}^-$	$HSO_{3(aq)}^-$
I	0.300	0	0
C	-x	+x	+x
E	0.300 - x	+x	+x

$$K_b = \frac{K_w}{K_{a2}} = \frac{1.0 \times 10^{-14}}{6.3 \times 10^{-8}} = 1.59 \times 10^{-7}$$

$$K_b = \frac{[OH_{(aq)}^-][HSO_{3(aq)}^-]}{[SO_{3(aq)}^{2-}]} = 1.59 \times 10^{-7} = \frac{x^2}{0.300 - x} \sim \frac{x^2}{0.300} \text{ since } 10^{-7} \ll 10^{-1}$$

$$[OH_{(aq)}^-] = 2.18 \times 10^{-4} \underline{M} \therefore pH = 14 + \log(2.18 \times 10^{-4} \underline{M}) = 10.34$$

c.) 0.300 L solution containing 0.500 M NaHSO₃ and 0.300 M Na₂SO₃



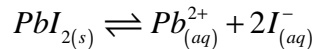
$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{SO}_{3(aq)}^{2-}]}{[\text{HSO}_{3(aq)}^-]} \right)$$

$$\text{pH} = -\log(6.3 \times 10^{-8}) + \log \left(\frac{0.300}{0.500} \right)$$

$$\text{pH} = 6.98$$

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

a.) 0.500 M PbI₂ with K_{sp} = 1.4 x 10⁻⁸

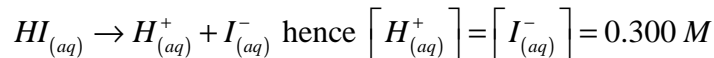


	$\text{Pb}_{(aq)}^{2+}$	$\text{I}_{(aq)}^-$
I	0	0
C	+s	+2s
E	+s	+2s

$$K_{sp} = 1.4 \times 10^{-8} = s(2s)^2 = 4s^3$$

$$s = 0.0015 \text{ M} \quad [\text{Pb}_{(aq)}^{2+}] = 0.0015 \text{ M} \quad \& \quad [\text{I}_{(aq)}^-] = 0.0030 \text{ M}$$

b.) 0.300 M HI



c.) 1 L solution containing 0.500 M PbI₂ and 0.300 M HI

	$\text{Pb}_{(aq)}^{2+}$	$\text{I}_{(aq)}^-$
I	0	0.300
C	+s	+2s
E	+s	0.300+2s

$$K_{sp} = 1.4 \times 10^{-8} = s(0.300 + 2s)^2$$

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

$$\text{ck assumption: } \frac{1.56 \times 10^{-7}}{0.300} \times 100\% = 5 \times 10^{-5} \%$$

$$[\text{Pb}_{(aq)}^{2+}] = 1.6 \times 10^{-7} \text{ M} \quad \& \quad [\text{I}_{(aq)}^-] = 0.300 \text{ M}$$

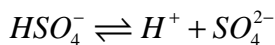
3. . Calculate the pOH and the concentrations of all species in a solution prepared by mixing equal volumes of 0.2 M HCl and 0.6 M H₂SO₄ (K_{a2} = 1.2 x 10⁻²).

we have two strong acids which completely dissociate:



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

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



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

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

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

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

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

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

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

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

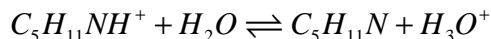
a.) $C_5H_{11}N$ (piperidine - $K_b = 1.3 \times 10^{-3}$) & HNO_3

When we add equal amounts of acid and base they react completely to produce



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.



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

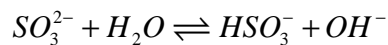
$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.3 \times 10^{-3}} = 7.7 \times 10^{-12}$$

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

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

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 $0.10 \text{ M } \text{SO}_3^{2-}$ in water.



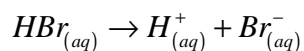
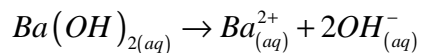
	SO_3^{2-}	OH^-	HSO_3^-
I	0.10	0	0
C	-x	+x	+x
E	$0.10 - x$	+x	+x

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

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

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

c.) $\text{Ba}(\text{OH})_2$ & HBr



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

