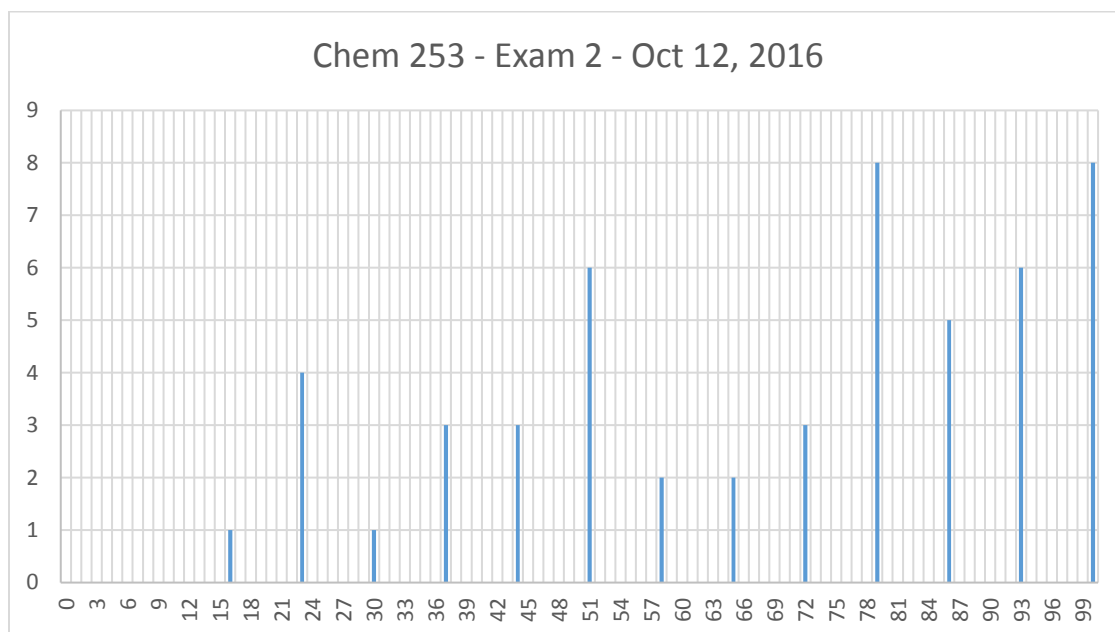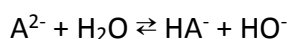Chem 253 Exam 2 – October 12, 2016

Chem 253 - Exam 2 - Oct 12, 2016

1] Calculate the pH of 0.100 F weak diprotic acid H_2A where $K_{a1} = 1.00 \times 10^{-4}$ and $K_{a2} = 1.0 \times 10^{-9}$.[1]

2] Using the constants from problem 1 what is K for the following? [2]

$$A^{2-} + H_2O \rightleftarrows HA^- + HO^-$$

3] What is the solubility of AgCl ($K_{sp} = 1.8 \times 10^{-10}$) in 0.100 M KCl? [3]

4] What is the solubility of Ag_2SO_4 ($K_{sp} = 1.5 \times 10^{-5}$). [4]

5] What is the MBE for 0.1 F H_3AsO_4? $K_{a1} = 5.8 \times 10^{-3}$, $K_{a2} = 1.10 \times 10^{-7}$, $K_{a3} = 3.2 \times 10^{-12}$ [5]

6] What is the CBE for a solution that is 0.1 M $MgCl_2$ and 0.2 M in $NaHCO_3$?   [6]

7] A sample weighing 3.000 g is a mixture of Fe_2O_3 (MW = 159.69) and Al_2O_3 (MW = 101.96). A heated stream of H_2 gas converts the Fe_2O_3 to Fe(s) (AW = 55.85) and $H_2O(g)$ (MW = 18.02). The Al_2O_3 does not react. The sample is weighed again at 2.500 g. What is the mass fraction of Fe_2O_3 in the original sample? [7]

8] Given that $K_a = 7.5 \times 10^{-6}$ for a weak acid where you describe its best buffer region?   [8]

9] Given that a weak diprotic acid H_2A has $K_{a1} = 1.00 \times 10^{-4}$ and $K_{a2} = 1.0 \times 10^{-9}$, calculate the pH of 0.100 F NaHA.   [9]

10] For that weak acid in problem 9 the fraction in the form of A^{2-} at pH 8.50 is best represented as which of the following?   [10]

a) $\alpha A^{2-} = [H^+]^2 / [H^+]^2 + [H^+]K_{a1} + K_{a1}K_{a2}$

b) $\alpha A^{2-} = K_{a1}K_{a2} / [H^+] + [H^+]K_{a1} + K_{a1}K_{a2}$

c) $\alpha A^{2-} = [H^+]K_{a2} / [H^+]^2 + [H^+]K_{a1} + K_{a1}K_{a2}$

d) $\alpha A^{2-} = K_{a1}K_{a2} / [H^+]^2 + [H^+]K_{a1} + K_{a1}K_{a2}$

e) $\alpha A^{2-} = K_{a1}K_{a2} / [H^+]^2 + [H^+]K_{a2} + K_{a1}K_{a2}$

A 50.00 mL sample of 0.100 M weak diprotic acid H_2A with $K_{a1} = 1.00 \times 10^{-4}$ and $K_{a2} = 1.0 \times 10^{-9}$ is titrated with 0.100 M NaOH. Note – there is 5.00 mmol H_2A in that sample.

11] What is the pH of the titration solution when 25.00 mL of 0.100 M NaOH (2.50 mmol) is added?   [11]

12] Which of the following allows calculation of the pH of the titration solution when 50.00 mL of 0.100 M NaOH (5.00 mmol) is added?   [12]

    a) $pH = \frac{1}{2}(pK_{a1} + pK_{a2})$

    b) $K_w/K_{a1} = x^2 / (0.050 - x)$

    c) $pH = pK_{a1}$

    d) $K_{a1} = x^2 / (0.050 - x)$

    e) $pH = pK_{a2}$

13] Which of the following allows calculation of the pH of the titration solution when 75.00 mL of 0.100 M NaOH (7.50 mmol) is added?   [13]

    a) $pH = \frac{1}{2}(pK_{a1} + pK_{a2})$

    b) $K_w/K_{a1} = x^2 / (0.050 - x)$

    c) $pH = pK_{a1}$

    d) $K_{a1} = x^2 / (0.067 - x)$

    e) $pH = pK_{a2}$

14] Which of the following allows calculation of the pH of the titration solution when 100.00 mL of 0.100 M NaOH (10.00 mmol) is added?   [14]

    a) $pH = \frac{1}{2}(pK_{a1} + pK_{a2})$
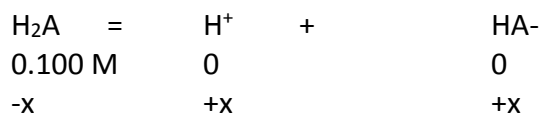
    b) $K_w/K_{a2} = x^2 / (0.033 - x)$

    c) $pH = pK_{a1}$

    d) $K_{a2} = x^2 / (0.033 - x)$

    e) $pH = pK_{a2}$

Answers

_____
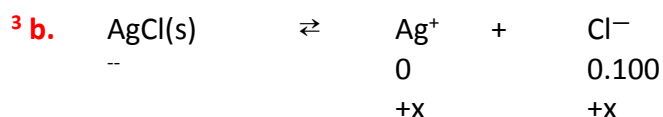
[1] **d.** $K_{a1} = 1.00 \times 10^{-4}$ >> K_{a2} is 1×10^{-9} so only K_{a1} is important.

H_2A	=	H^+	+	$HA-$
0.100 M		0		0
-x		+x		+x

$K_{a1} = 1.00 \times 10^{-4} = x^2/0.100-x = x^2/0.100$ x = 3.16e-3   pH = 2.500

[2] **a.** This is K_{b1} $K_{b1} = K_w/K_{a2} = 1.00e-14/1.0e-9 = 1.0e-5$

[3] **b.**

$AgCl(s)$	$\rightleftarrows$	Ag^+	+	Cl^-
--		0		0.100
		+x		+x

Ksp = 1.8e-10 = x(0.10+x) $\cong$ x(0.100) x = 1.8e-9 M

[4] **a.**

Ag_2SO_4	$\rightleftarrows$	$2Ag^+$	+	SO_4^{2-}
-		+2x		+x
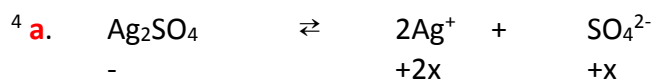
$(2x)^2x = 4x^3 = 1.5e-5$ x = 1.6e-2 M

[5] **e.**   0.1 F H_3AsO_4 = $[H_3AsO_4]$ + $[H_2AsO_4^-]$ + $[HAsO_4^{2-}]$ + $[AsO_4^{3-}]$

[6] **c.**   Note $HCO_3^- \rightleftarrows H^+ + CO_3^{2-}$ and $HCO_3^- + H_2O \rightleftarrows H_2CO_3 + OH^-$

$2[Mg^{2+}] + [Na^+] + [H^+] = [HCO_3^-] + 2[CO_3^{2-}] + [OH^-] + [Cl^-]$

[7] **e.** 0.500 g O (mol O/15.999 g) (mol Fe_2O_3/ 3 mol O) (159.69 g / mol Fe_2O_3) (100 / 3.000) = 55.4%

[8] **c.** $K_a = 7.5 \times 10^{-6}$ pKa = 5.12   general rule buffer region +/- 1 pH with pKa so 4.12 to 6.12
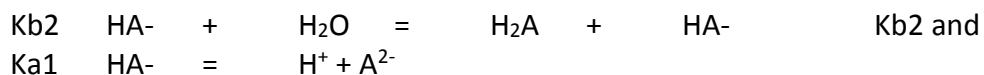
[9] **b.** High formal concentrations mean that pH = ½(pKa1 + pKa2) works well.

pH = ½(4.000 + 9.00) = 6.50

10 d. $\alpha^{2-} = K_{a1}K_{a2} / [H^+]^2 + [H^+]K_{a1} + K_{a1}K_{a2}$

11 a.   ½ first eq pt.   of $[H_2A] = [HA^-]$
               $K_{a1} = [H^+]$ or pH = pKa = 4.00   1st buffer region

12 a. This is the first eq pt. where initially all HA- or [HA-] = 5.00 mmol/ 100.00 mL = 0.050 M

Kb2   HA-   + H_2O = H_2A +   HA-   Kb2 and
Ka1   HA-   = $H^+ + A^{2-}$
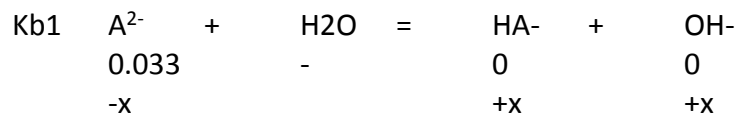
Amphoteric Species   pH = ½ (pKa1 + pKa2) = 6.50

13 e. 2.50 mmol past 1st eq pt. So we have 2.50 mmol HA- and 2.50 mmol of A^{2-} 2nd buffer region

$K_{a2} = [H^+][A^{2-}] / [HA^-]$ and $[A^{2-}] = [HA^-]$ so   pH = pKa2

14 b. 2nd eq pt. so $[A^{2-}]$ = 5.00 mmol / 150.00 = 0.0333 M

Kb1 A^{2-} +   H2O   =   HA-   +   OH-
         0.033   -   0   0
         -x   +x   +x

$Kb1 = K_w/K_{a2} = x^2 / (0.033 - x) = 1.00e-14/1.00e-9 = 1.00e-5$
X = 5.745e-4   pH = 10.76