

CHEM 1423
Chapters 19
Homework Solutions

TEXTBOOK HOMEWORK

19.12 $C_6H_5ONa \rightleftharpoons Na^+ + C_6H_5O^-(1.3 M)$, $C_6H_5OH(1.2 M)$, $pK_a = 10.00$

$$pH = pK_a + \log\left(\frac{[C_6H_5O^-]}{[C_6H_5OH]}\right) = 10.00 + \log\left(\frac{1.3}{1.2}\right) = 10.03$$

19.15 $CH_3NH_3Cl \rightleftharpoons Cl^- + CH_3NH_3^+(0.60 M)$, $CH_3NH_2(0.50 M)$

$$pK_b(CH_3NH_2) = 3.35 \quad \parallel \quad pK_a(CH_3NH_3^+) = 14. - 3.35 = 10.65$$

$$pH = pK_a + \log\left(\frac{[CH_3NH_2]}{[CH_3NH_3^+]}\right) = 10.65 + \log\left(\frac{0.50}{0.60}\right) = 10.57$$

19.16 $HPr \rightleftharpoons H^+ + Pr^-$, $K_a = 1.3 \times 10^{-5}$ \parallel $pK_a = 4.89$, $pH = 5.44$

$$pH = pK_a + \log\left(\frac{[Pr^-]}{[HPr]}\right)$$

$$5.44 = 4.89 + \log\left(\frac{[Pr^-]}{[HPr]}\right)$$

$$\log\left(\frac{[Pr^-]}{[HPr]}\right) = 5.44 - 4.89 = +0.55$$

$$\frac{[Pr^-]}{[HPr]} = 10^{+0.55} = 3.55$$

NOTE: Had the question asked instead for the ratio, $[HPr]/[Pr^-]$ (i.e. the ratio of acid/base), you would work the problem in the same way, and then invert at the end;

$$\frac{[HPr]}{[Pr^-]} = \frac{1}{[Pr^-]/[HPr]} = \frac{1}{3.55} = 0.28$$

19.18 We will work this in 2 parts: (1) We'll use initial data to determine pK_a , (2) We'll then determine the concentrations after the addition of NaOH, from which we'll get the new pH.

Note: This problem is very similar to #19.19

(1) $[HA] = 0.2000 \text{ M}$, $[A^-] = 0.1500 \text{ M}$, $\text{pH} = 3.35$

$$\text{pH} = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

$$3.35 = pK_a + \log\left(\frac{0.150}{0.20}\right)$$

$$3.35 = pK_a - 0.12$$

$$pK_a = 3.35 + 0.12 = 3.47$$

(2) Initial n_{HA} and n_{A^-}

$$n_{HA} = 0.20 \text{ mol} / \text{L} \times 0.50 \text{ L} = 0.100 \text{ mol HA}$$

$$n_{A^-} = 0.150 \text{ mol} / \text{L} \times 0.50 \text{ L} = 0.0750 \text{ mol } A^-$$

Now let's determine concentrations after adding 0.0015 mol OH^- .

	HA	OH^-	A^-
n(Init)	0.100	0.0015	0.0750
n(Chg)	-0.0015	-0.0015	+0.0015
n(Fin)	0.0985	0	0.0765
Conc*	0.1970	0	0.1530

*Conc. was determined by dividing n(Fin) by 0.50 L

$$\text{pH} = pK_a + \log\left(\frac{[A^-]}{[HA]}\right) = 3.47 + \log\left(\frac{0.153}{0.197}\right) = 3.36$$

Note: One would get the same result using the values of n(Fin) rather than the concentration

19.19 We will work this in 2 parts: (1) We'll use initial data to determine pK_a , (2) We'll then determine the concentrations after the addition of HCl, from which we'll get the new pH.

Note: This problem is very similar to #19.18

(1) $[B] = 0.400 \text{ M}$, $[BH^+] = 0.250 \text{ M}$, $pH = 8.88$

$$pH = pK_a + \log\left(\frac{[B]}{[BH^+]}\right)$$

$$8.88 = pK_a + \log\left(\frac{0.400}{0.250}\right)$$

$$8.88 = pK_a + 0.20$$

$$pK_a = 8.88 - 0.20 = 8.68$$

(2) Initial n_B and n_{BH^+}

$$n_B = 0.40 \text{ mol} / \text{L} \times 0.250 \text{ L} = 0.100 \text{ mol } B$$

$$n_{BH^+} = 0.250 \text{ mol} / \text{L} \times 0.250 \text{ L} = 0.0625 \text{ mol } A^-$$

Now let's determine concentrations after adding 0.0020 mol H^+ .

	B	H^+	BH^+
n(Init)	0.100	0.0020	0.0625
n(Chg)	-0.0020	-0.0020	+0.0020
n(Fin)	0.0980	0	0.0645
Conc*	0.392	0	0.258

*Conc. was determined by dividing n(Fin) by 0.25 L

$$pH = 8.68 + \log\left(\frac{[B]}{[BH^+]}\right) = 8.68 + \log\left(\frac{0.392}{0.258}\right) = 8.86$$

Note: One would get the same result using the values of n(Fin) rather than the concentration

19.42(C) Triethyl Amine = Et_3N , $K_b = 5.2 \times 10^{-4}$. Add 15.00 mL of 0.100 M HCl to 20.00 mL of 0.100 M Et_3N .

Note: The addition of the strong acid, HCl, will convert some of the Et_3N into Et_3NH^+ via the reaction $\text{Et}_3\text{N} + \text{H}^+ \rightleftharpoons \text{Et}_3\text{NH}^+$. After the addition of the H^+ , we will have a buffer containing $[\text{Et}_3\text{N}]$ and $[\text{Et}_3\text{NH}^+]$, and will use the Henderson-Hasselbalch to determine the pH.

(1) Determine the $\text{p}K_a$ of Et_3NH^+ .

$$K_b = 5.2 \times 10^{-4}$$

$$\text{p}K_b = -\log(5.2 \times 10^{-4}) = 3.28$$

$$\text{p}K_a = 14. - 3.28 = 10.72$$

(2) Determine initial amounts of Et_3N and H^+

$$n_{\text{Et}_3\text{N}} = 0.100 \text{ mol} / \text{L} \times 0.020 \text{ L} = 0.0020 \text{ mol } \text{Et}_3\text{N}$$

$$n_{\text{H}^+} = 0.100 \text{ mol} / \text{L} \times 0.0150 \text{ L} = 0.00150 \text{ mol } \text{H}^+$$

(3) Determine Final Concentrations and pH

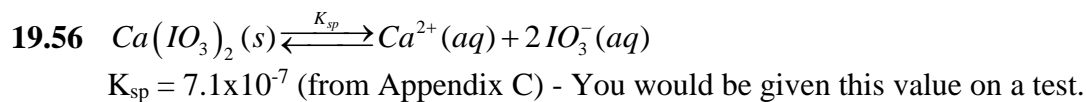
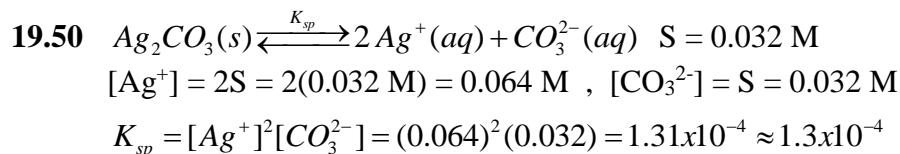
	Et_3N	H^+	Et_3NH^+
n(Init)	0.0020	0.00150	0
n(Chg)	-0.00150	-0.00150	+0.00150
n(Fin)	0.00050	0	0.00150
Conc*	0.01429	0	0.04286

*Concentration determined by dividing by 35 mL = 0.035 M

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{Et}_3\text{N}]}{[\text{Et}_3\text{NH}^+]}\right) = 10.72 + \log\left(\frac{0.01429}{0.04286}\right)$$

$$\text{pH} = 10.24$$

Note: One would get the same result using the values of n(Fin) rather than the concentration



Preliminary: For review, let's calculate the solubility of $Ca(IO_3)_2$ in pure water.

$$[Ca^{2+}] = S, IO_3^- = 2S$$

$$K_{sp} = 7.1 \times 10^{-7} = [Ca^{2+}][IO_3^-]^2 = S(2S)^2 = 4S^3$$

$$S^3 = \frac{7.1 \times 10^{-7}}{4} = 1.78 \times 10^{-7}$$

$$S = (1.78 \times 10^{-7})^{1/3} = 5.62 \times 10^{-3}\text{ M}$$

Note: If we asked for $[IO_3^-]$ in pure water, it would be $2S = 1.12 \times 10^{-2}\text{ M}$

(a) Solubility in 0.06 M $Ca(NO_3)_2$ [\parallel $Ca^{2+}(0.06\text{ M}) + 2NO_3^-$]

$$[Ca^{2+}] = 0.06 + S \approx 0.06\text{ M}, IO_3^- = 2S$$

$$K_{sp} = 7.1 \times 10^{-7} = [Ca^{2+}][IO_3^-]^2 = (0.06)(2S)^2 = 0.24S^2$$

$$S^2 = \frac{7.1 \times 10^{-7}}{0.24} = 2.96 \times 10^{-6}$$

$$S = (2.96 \times 10^{-6})^{1/2} = 1.72 \times 10^{-3}\text{ M}$$

(b) Solubility in 0.06 M $NaIO_3$ [\parallel $Na^+ + IO_3^-(0.06\text{ M})$]

$$[Ca^{2+}] = S, IO_3^- = 0.06 + 2S \approx 0.06$$

$$K_{sp} = 7.1 \times 10^{-7} = [Ca^{2+}][IO_3^-]^2 = (S)(0.06)^2 = (3.6 \times 10^{-3})S$$

$$S^2 = \frac{7.1 \times 10^{-7}}{3.6 \times 10^{-3}} = 1.97 \times 10^{-4} \approx 2 \times 10^{-4}$$