Chemistry 102 – Exam 2 Name______________________
12 April 2010

Show all work for credit. Answers without supporting work will receive zero (0) points. Give all numerical answers with the correct number of significant figures and the correct units.

1. 10 points – Calculate the ***K_a* of the conjugate acid** of a monoprotic base if a 0.24 M solution of the base has a pH of 11.87.

$B + H_2O \leftrightharpoons HB^+ + OH^-$

	B	HB^+	OH^-
I	0.24	0	1.0×10^{-7}
C	-x	+x	+x
E	0.24 - x	x	1.0×10^{-7} + x

$$\left[OH^-\right] = 10^{-pOH} = 10^{-(14.00-pH)} = 10^{-(14.00-11.87)} = 7.413\times10^{-3} \approx \left[HB^+\right]$$

$$K_b = \frac{\left[OH^-\right]\left[HB^+\right]}{[B]} \approx \frac{\left(7.413\times10^{-3}\right)^2}{0.24} = 2.3\times10^{-4}$$

$$K_a = \frac{K_w}{K_b} = \frac{1.0\times10^{-14}}{2.3\times10^{-4}} = 4.3\times10^{-11}$$

2. 20 points – Modified Textbook Problem 17.89 – Ammonia, NH_3, is a base that ionizes to give NH_4^+ and OH^- ($K_b = 1.8\times10^{-5}$). You add magnesium sulfate to an ammonia solution. The concentration of Mg^{2+} ion when magnesium hydroxide, $Mg(OH)_2$, just begins to precipitate from 0.10 M NH_3 is 1.0×10^{-5} M. Calculate the value of ***K_{sp}* for magnesium hydroxide**.

$NH_3 + H_2O \leftrightharpoons NH_4^+ + OH^-$ $K_b = 1.8\times10^{-5}$

	NH_3	NH_4^+	OH^-
I	0.10	0	1.0×10^{-7}
C	-x	+x	+x
E	0.10 - x	x	1.0×10^{-7} + x

$$K_b = \frac{\left[NH_4^+\right]\left[OH^-\right]}{\left[NH_3\right]} = \frac{(x)\left(1.0\times10^{-7}+x\right)}{(0.10-x)} \approx \frac{x^2}{0.10} \therefore x = \sqrt{\left(1.8\times10^{-5}\right)(0.10)} = 0.0013 \approx \left[OH^-\right]$$

$$K_{sp} = \left[Mg^{2+}\right]\left[OH^-\right]^2 = \left(1.0\times10^{-5}\right)(0.0013)^2 = 1.8\times10^{-11}$$

3. 20 points – Does a ***precipitate*** form when 0.0456 g of Iron(III) Nitrate is added to 250.0 mL of a solution that is 2.00×10^{-5} M Sodium Hydroxide and 1.00×10^{-4} M Sodium Cyanide? Show all work that leads you to your answer. No points for just a "yes" or "no" answer.

$$Fe(OH)_3 \rightleftharpoons Fe^{3+} + 3\ OH^- \quad K_{sp} = 2.6\text{x}10^{-39}$$
$$Fe^{3+} + 6\ CN^- \rightleftharpoons Fe(CN)_6^{3-} \quad K_f = 1.0\text{x}10^{31}$$

$$Fe(OH)_3 + 6\ CN^- \rightleftharpoons Fe(CN)_6^{3-} + 3\ OH^- \quad K = 2.6\text{x}10^{-8}$$

Because the Fe^{3+} is not the limiting reactant with the CN^- we have to calculate the amount of Fe^{3+} left after combining with the CN^-.

$$\left[Fe^{3+}\right]_{initial} = \frac{0.0456\text{ g Fe}(NO_3)_3}{250.0\text{ mL sol'n}} \times \frac{1\text{ mL}}{10^{-3}\text{ L}} \times \frac{1\text{ mol Fe}(NO_3)_3}{241.861\text{ g Fe}(NO_3)_3} \times \frac{1\text{ mol Fe}^{3+}}{1\text{ mol Fe}(NO_3)_3} \times \frac{1\text{ mol Fe}(CN)_6^{3-}}{1\text{ mol Fe}^{3+}} = 0.000754\text{ M}$$

$$\left[Fe(CN)_6^{3-}\right] = \frac{1.00\times10^{-4}\text{ mol CN}^-}{\text{L}} \times \frac{1\text{ mol Fe}(CN)_6^{3-}}{6\text{ mol CN}^-} = 1.7\times10^{-5}\text{ M}$$

$$\left[Fe^{3+}\right]_{free} = \left[Fe^{3+}\right]_{initial} - \left[Fe(CN)_6^{3-}\right] = 0.000754 - 1.7\times10^{-5} = 7.4\times10^{-4}$$

$$Q = \left[Fe^{3+}\right]\left[OH^-\right]^3 = \left(7.4\times10^{-4}\right)\left(2.00\times10^{-5}\right)^3 = 5.9\times10^{-18} > K_{sp}$$

So a precipitate does form.

4. 20 points – Calculate the ***concentration of free tin(IV)*** ions in a solution that initially contains 0.050 M tin(IV) nitrate and 0.750 M sodium fluoride.

$$Sn^{4+} + 6\ F^- \rightleftharpoons SnF_6^{2-} \quad K_f = 1\text{x}10^{25}$$

	Sn^{4+}	F^-	SnF_6^{2-}
I	0.050	0.750	0
Completion	-0.050	-0.300	+0.050
C	+x	+6x	-x
E	x	0.450 + 6x	0.050 – x

$$K_f = \frac{\left[SnF_6^{2-}\right]}{\left[Sn^{4+}\right]\left[F^-\right]^6} = \frac{(0.050 - x)}{x(0.450 + 6x)^6} \approx \frac{0.050}{x(0.450)^6} = 1\times10^{25}$$

$$x = \frac{0.050}{\left(1\times10^{25}\right)(0.450)^6} = 6\times10^{-25} \approx \left[Sn^{4+}\right]$$

5. 15 points – Calculate the ***molar solubility of lead(II) chromate*** in solution that contains 0.015 M sodium chromate.

$$PbCrO_{4(s)} \leftrightharpoons Pb^{2+}_{(aq)} + CrO_4^{2-}_{(aq)} \qquad K_{sp} = 2.8\times10^{-13}$$

	Pb^{2+}	CrO_4^{2-}
I	0	0.015
C	+x	+x
E	x	0.015+x

$$K_{sp} = \left[Pb^{2+}\right]\left[CrO_4^{2-}\right] = x(0.015+x) \approx 0.015x \approx 2.8\times10^{-13}$$

$$x = 1.9\times10^{-11}\text{ M}$$

6. 15 points – Calculate the ***pH*** of a solution that is obtained by combining 150.0 mL of 0.251 M Cyanic acid with 100.0 mL of 0.076 M Calcium Hydroxide.

$$\text{mmol HCNO} = 150.0\text{ mL} \times \frac{0.251\text{ mol HCNO}}{\text{L}} = 37.65\text{ mmol HCNO}$$

$$\text{mmol OH}^- = 100.0\text{ mL} \times \frac{0.076\text{ mol Ca(OH)}_2}{\text{L}} \times \frac{2\text{ mol OH}^-}{1\text{ mol Ca}(OH)_2} = 15.2\text{ mmol OH}^-$$

$HCNO$	$+\ OH^-$	$\leftrightharpoons H_2O\ +$	CNO^-
37.65	15.2		0
-15.2	-15.2		+15.2
22.4	0		15.2

This is a buffer

$$pH = pK_a + \log\frac{\left[CNO^-\right]}{[HCNO]} = -\log\left(3.5\times10^{-4}\right) + \log\frac{15.2}{22.4} = 3.29$$