1. Calculate the K_a of a monoprotic acid if a 0.153 M solution of the acid has a pH of 6.649.

$$\left[H_3O^+\right]_{eq} = 10^{-pH} = 10^{-6.649} = 2.24\times10^{-7}\text{ M}$$

	$HA + H_2O \leftrightharpoons$	H_3O^+	$+ A^-$
	0.153	10^{-7}	0
	-x	+x	+x
	0.153-x	$x+10^{-7}$	x

$$x + 1.0\times10^{-7} = \left[H_3O^+\right]_{eq} = 2.24\times10^{-7} \therefore x = 1.24\times10^{-7}$$

$$0.153 - x \approx 0.151 = [HA]_{eq}$$

$$x = \left[A^-\right]_{eq} = 1.24\times10^{-7}$$

$$K_a = \frac{\left[A^-\right]_{eq}\left[H_3O^+\right]_{eq}}{[HA]_{eq}} = \frac{\left(1.24\times10^{-7}\right)\left(2.24\times10^{-7}\right)}{0.151} = 1.84\times10^{-13}$$

2. 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.

From Ammonia we get the OH^- concentration

	$NH_3 + H_2O \leftrightharpoons$	OH^-	$+ NH_4^+$
	0.10 M	10^{-7}	0
	-x	+x	+x
	0.10-x	10^{-7}+x	x

$$K_b = 1.8\times10^{-5} = \frac{\left[OH^-\right]\left[NH_4^+\right]}{\left[NH_3\right]} = \frac{\left(10^{-7}+x\right)(x)}{(0.10-x)} \approx \frac{x^2}{0.10}$$

$$x = \left[OH^-\right] = \sqrt{(0.10)\left(1.8\times10^{-5}\right)} = 1.3\times10^{-3}$$

assumptions are valid

Find K_{sp} from above information:

$$K_{sp} = \left[Mg^{2+}\right]\left[OH^-\right]^2 = \left(1.0\times10^{-5}\right)\left(1.3\times10^{-3}\right)^2 = 1.8\times10^{-11}$$

3. Qualitative analysis.

a. Multiple Choice (circle the correct answer)

Which of the following ions forms a precipitate when aqueous Chloride ion is added?

Li^{+} NH_4^{+} Ag^{+} Cu^{2+}

Which of the following ions gives a persistent bright yellow color in a flame test?

Li^{+} NH_4^{+} Na^{+} K^{+}

Which of the following ions forms a red-maroon precipitate with $K_4Fe(CN)_6$?

Li^{+} Cr^{3+} Cu^{2+} Cu^{+}

Which of the following ions can be precipitated with aqueous ammonia?

Fe^{3+} Li^{+} Cu^{2+} Na^{+}

Which of the following can be used as a confirmation of the presence of Cu^{2+}?

$HCl_{(aq)}$ $K_4Fe(CN)_{6\,(aq)}$ $HNO_{3\,(aq)}$ $NH_{3\,(aq)}$

b. Separations

A solution contains Cu^{2+} and Cr^{3+} ions. Give one (***and only one***) test that will separate these two ions. State which ion is in which phase (supernatant or precipitate). Giving more than one test will result in zero points.

Separation: Add aqueous ammonia

Cu^{2+} will be in the supernatant
Cr^{3+} will be in the precipitate as a black solid

4. A solution is a mixture of 0.15 M Silver Nitrate and 0.26 M Barium Nitrate. A solution of Sodium Oxalate is added dropwise.

a. Which ion will precipitate first? (Show work, no points for lucky guesses)

Calculate the concentration of Oxalate needed to precipitate both. The one requiring the lower concentration will precipitate first.

$$K_{sp} = \left[Ag^{+}\right]^{2}\left[C_2O_4^{2-}\right] = 3.5\times10^{-11} = (0.15)^{2}(x) \Rightarrow x = \left[C_2O_4^{2-}\right] = 1.6\times10^{-9}\text{ M}$$
$$K_{sp} = \left[Ba^{2+}\right]\left[C_2O_4^{2-}\right] = 1.6\times10^{-7} = (0.26)(x) \Rightarrow x = \left[C_2O_4^{2-}\right] = 6.1\times10^{-7}\text{ M}$$

Silver will precipitate first.

b. What percent of the ion that precipitates first will remain in solution when the second metal ion starts to precipitate?

Calculate the Silver ion concentration when the Barium ion starts to precipitate:

$$K_{sp} = \left[Ag^{+}\right]^{2}\left[C_2O_4^{2-}\right] = 3.5\times10^{-11} \Rightarrow \left[Ag^{+}\right] = \sqrt{\frac{3.5\times10^{-11}}{6.1\times10^{-7}}} = 7.5\times10^{-3}\text{ M}$$

$$\% = \frac{7.5\times10^{-3}}{0.15}\times100 = 5.0\%$$

5. Calculate the ***pH*** of a solution that is obtained by combining 150.0 mL of 0.151 M Hydrazine with 200.0 mL of 0.096 M Nitric acid.

$$n_{N_2H_4} = 150.0\text{ mL} \times \frac{0.151\text{ mol N}_2\text{H}_4}{L} = 22.7\text{ mmol N}_2\text{H}_4$$

$$n_{H^+} = 200.0\text{ mL} \times \frac{0.096\text{ mol HNO}_3}{L} \times \frac{1\text{ mol H}^+}{1\text{ mol HNO}_3} = 19.\text{ mmol H}^+$$

Chem 101 first because we have a strong acid so the reaction goes to completion

N_2H_4	+	H^+	$\rightarrow$	$N_2H_3^+$	
22.7 mmol		19. mmol		0 mmol	H^+ is limiting
-19.		-19.		+19.	
3. mmol		0		19. mmol	It's a buffer!!!

$$pH = pK_a + \log\frac{n_{N_2H_4}}{n_{N_2H_5^+}} = -\log\left(\frac{1.0\times10^{-14}}{1.3\times10^{-6}}\right) + \log\frac{3.}{19.} = 7.3$$

6. Does a precipitate form when 0.0456 g of Nickel(II) Nitrate is added to 250.0 mL of a solution that is 1.00 M Sodium Hydroxide and 2.00 M Sodium Cyanide? Show all work that leads you to your answer. No points for just a "yes" or "no" answer.

Calculate $[Ni^{2+}]$ before and after the complex ion $Ni(CN)_4^{2-}$ forms ($K_f = 1.1\times10^{31}$)

$$\left[Ni^{2+}\right]_i = \frac{0.465 \text{ g Ni}\left(\text{NO}_3\right)_2}{250.0 \text{ mL}} \times \frac{1 \text{ mol Ni}\left(\text{NO}_3\right)_2}{182.7032 \text{ g Ni}\left(\text{NO}_3\right)_2} \times \frac{1 \text{ mL}}{10^{-3} \text{ L}} = 9.98\times10^{-4} \text{ M}$$

Ni^{2+}	+	$4\ CN^-$	⇋	$Ni(CN)_4^{2-}$	
9.98×10^{-4}		2.00		0	go to completion first
-9.98×10^{-4}		$-4(9.98\times10^{-4})$		$+9.98\times10^{-4}$	
0		1.996		9.98×10^{-4}	go back to equilibrium
+x		+4x		-x	
x		1.996 + 4x		9.98×10^{-4}-x	

$$K_f = 1.1\times10^{31} = \frac{\left[Ni(CN)_4^{2-}\right]_{eq}}{\left[Ni^{2+}\right]_{eq}\left[CN^-\right]_{eq}^4} = \frac{\left(9.98\times10^{-4} - x\right)}{(x)(1.996+4x)^4} \approx \frac{9.98\times10^{-4}}{x(1.996)^4} \Rightarrow x = 5.8\times10^{-36} = \left[Ni^{2+}\right]$$

Find Q_{sp} for the Nickel(II) ion and the hydroxide ion:

$$Q_{sp} = \left[Ni^{2+}\right]\left[OH^-\right]^2 = \left(5.8\times10^{-36}\right)(1.00)^2 = 5.8\times10^{-36} << K_{sp}$$

Therefore, we would not expect a precipitate to form.