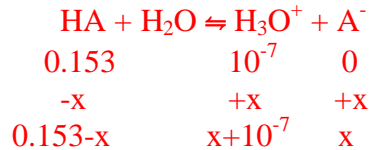


1. Calculate the  $K_a$  of a monoprotic acid if a 0.153 M solution of the acid has a pH of 6.649.

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



$$x + 1.0 \times 10^{-7} = [H_3O^+]_{eq} = 2.24 \times 10^{-7} \therefore x = 1.24 \times 10^{-7}$$

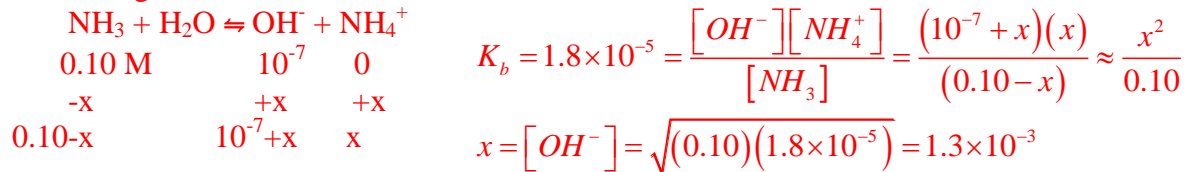
$$0.153 - x \approx 0.151 = [HA]_{eq}$$

$$x = [A^-]_{eq} = 1.24 \times 10^{-7}$$

$$K_a = \frac{[A^-]_{eq} [H_3O^+]_{eq}}{[HA]_{eq}} = \frac{(1.24 \times 10^{-7})(2.24 \times 10^{-7})}{0.151} = 1.84 \times 10^{-13}$$

2. Ammonia,  $\text{NH}_3$ , is a base that ionizes to give  $\text{NH}_4^+$  and  $\text{OH}^-$  ( $K_b = 1.8 \times 10^{-5}$ ). You add magnesium sulfate to an ammonia solution. The concentration of  $\text{Mg}^{2+}$  ion when magnesium hydroxide,  $\text{Mg}(\text{OH})_2$ , just begins to precipitate from 0.10 M  $\text{NH}_3$  is  $1.0 \times 10^{-5}$  M. Calculate the value of  $K_{sp}$  for magnesium hydroxide.

From Ammonia we get the  $\text{OH}^-$  concentration



Find  $K_{sp}$  from above information:

assumptions are valid

$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2 = (1.0 \times 10^{-5})(1.3 \times 10^{-3})^2 = 1.8 \times 10^{-11}$$

3. Qualitative analysis.

a. Multiple Choice (circle the correct answer)

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



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



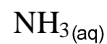
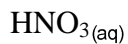
Which of the following ions forms a red-maroon precipitate with  $\text{K}_4\text{Fe}(\text{CN})_6$ ?



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



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



b. Separations

A solution contains  $\text{Cu}^{2+}$  and  $\text{Cr}^{3+}$  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

$\text{Cu}^{2+}$  will be in the supernatant

$\text{Cr}^{3+}$  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.

$$K_{sp} = [\text{Ag}^+]^2 [\text{C}_2\text{O}_4^{2-}] = 3.5 \times 10^{-11} = (0.15)^2 (x) \Rightarrow x = [\text{C}_2\text{O}_4^{2-}] = 1.6 \times 10^{-9} \text{ M}$$

$$K_{sp} = [\text{Ba}^{2+}] [\text{C}_2\text{O}_4^{2-}] = 1.6 \times 10^{-7} = (0.26)(x) \Rightarrow x = [\text{C}_2\text{O}_4^{2-}] = 6.1 \times 10^{-7} \text{ M}$$

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:

$$K_{sp} = [\text{Ag}^+]^2 [\text{C}_2\text{O}_4^{2-}] = 3.5 \times 10^{-11} \Rightarrow [\text{Ag}^+] = \sqrt{\frac{3.5 \times 10^{-11}}{6.1 \times 10^{-7}}} = 7.5 \times 10^{-3} \text{ M}$$

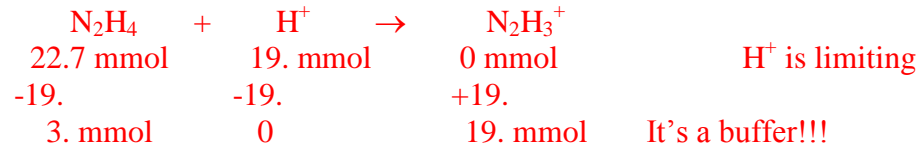
$$\% = \frac{7.5 \times 10^{-3}}{0.15} \times 100 = 5.0\%$$

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.

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

$$n_{H^+} = 200.0 \text{ mL} \times \frac{0.096 \text{ mol } HNO_3}{L} \times \frac{1 \text{ mol } H^+}{1 \text{ mol } HNO_3} = 19. \text{ mmol } H^+$$

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



$$pH = pK_a + \log \frac{n_{N_2H_4}}{n_{N_2H_3^+}} = -\log \left( \frac{1.0 \times 10^{-14}}{1.3 \times 10^{-6}} \right) + \log \frac{3.}{19.} = 7.3$$

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 ( $K_f = 1.1 \times 10^{31}$ )

$$[Ni^{2+}]_i = \frac{0.0456 \text{ g Ni(NO}_3)_2}{250.0 \text{ mL}} \times \frac{1 \text{ mol Ni(NO}_3)_2}{182.7032 \text{ g Ni(NO}_3)_2} \times \frac{1 \text{ mL}}{10^{-3} \text{ L}} = 9.98 \times 10^{-4} \text{ M}$$

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

$$K_f = 1.1 \times 10^{31} = \frac{[Ni(CN)_4^{2-}]_{eq}}{[Ni^{2+}]_{eq} [CN^-]_{eq}^4} = \frac{(9.98 \times 10^{-4} - x)}{(x)(1.996 + 4x)^4} \approx \frac{9.98 \times 10^{-4}}{x(1.996)^4} \Rightarrow x = 5.8 \times 10^{-36} = [Ni^{2+}]$$

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

$$Q_{sp} = [Ni^{2+}][OH^-]^2 = (5.8 \times 10^{-36})(1.00)^2 = 5.8 \times 10^{-36} \ll K_{sp}$$

Therefore, we would not expect a precipitate to form.