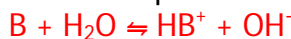


Chemistry 102 - Exam 2 Name _____
12 April 2010

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.



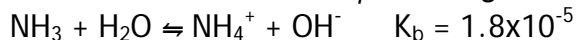
	B	HB ⁺	OH ⁻
I	0.24	0	1.0x10 ⁻⁷
C	-x	+x	+x
E	0.24 - x	x	1.0x10 ⁻⁷ + x

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

$$K_b = \frac{[OH^-][HB^+]}{[B]} \approx \frac{(7.413 \times 10^{-3})^2}{0.24} = 2.3 \times 10^{-4}$$

$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{2.3 \times 10^{-4}} = 4.3 \times 10^{-11}$$

2. 20 points - Modified Textbook Problem 17.89 - Ammonia, NH₃, is a base that ionizes to give NH₄⁺ and OH⁻ ($K_b = 1.8 \times 10^{-5}$). You add magnesium sulfate to an ammonia solution. The concentration of Mg²⁺ ion when magnesium hydroxide, Mg(OH)₂, just begins to precipitate from 0.10 M NH₃ is 1.0x10⁻⁵ M. Calculate the value of K_{sp} for magnesium hydroxide.

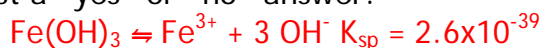


	NH ₃	NH ₄ ⁺	OH ⁻
I	0.10	0	1.0x10 ⁻⁷
C	-x	+x	+x
E	0.10 - x	x	1.0x10 ⁻⁷ + x

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{(x)(1.0 \times 10^{-7} + x)}{(0.10 - x)} \approx \frac{x^2}{0.10} \therefore x = \sqrt{(1.8 \times 10^{-5})(0.10)} = 0.0013 \approx [OH^-]$$

$$K_{sp} = [Mg^{2+}][OH^-]^2 = (1.0 \times 10^{-5})(0.0013)^2 = 1.8 \times 10^{-11}$$

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.



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^- .

$$[\text{Fe}^{3+}]_{\text{initial}} = \frac{0.0456 \text{ g Fe(NO}_3)_3}{250.0 \text{ mL sol'n}} \times \frac{1 \text{ mL}}{10^{-3} \text{ L}} \times \frac{1 \text{ mol Fe(NO}_3)_3}{241.861 \text{ g Fe(NO}_3)_3} \times \frac{1 \text{ mol Fe}^{3+}}{1 \text{ mol Fe(NO}_3)_3} \times \frac{1 \text{ mol Fe(CN)}_6^{3-}}{1 \text{ mol Fe}^{3+}} = 0.000754 \text{ M}$$

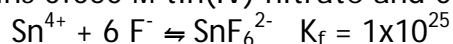
$$[\text{Fe(CN)}_6^{3-}] = \frac{1.00 \times 10^{-4} \text{ mol CN}^-}{\text{L}} \times \frac{1 \text{ mol Fe(CN)}_6^{3-}}{6 \text{ mol CN}^-} = 1.7 \times 10^{-5} \text{ M}$$

$$[\text{Fe}^{3+}]_{\text{free}} = [\text{Fe}^{3+}]_{\text{initial}} - [\text{Fe(CN)}_6^{3-}] = 0.000754 - 1.7 \times 10^{-5} = 7.4 \times 10^{-4}$$

$$Q = [\text{Fe}^{3+}][\text{OH}^-]^3 = (7.4 \times 10^{-4})(2.00 \times 10^{-5})^3 = 5.9 \times 10^{-18} > K_{sp}$$

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.

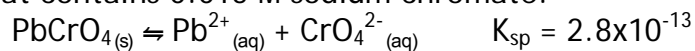


	Sn^{4+}	F^-	SnF_6^{2-}
I	0.050	0.750	0
Completion	-0.050	-0.300	+0.050
C	+x	+6x	-x
E	x	0.450 + 6x	0.050 - x

$$K_f = \frac{[\text{SnF}_6^{2-}]}{[\text{Sn}^{4+}][\text{F}^-]^6} = \frac{(0.050 - x)}{x(0.450 + 6x)^6} \approx \frac{0.050}{x(0.450)^6} = 1 \times 10^{25}$$

$$x = \frac{0.050}{(1 \times 10^{25})(0.450)^6} = 6 \times 10^{-25} \approx [\text{Sn}^{4+}]$$

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



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

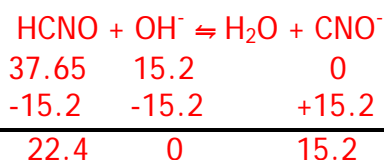
$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{CrO}_4^{2-}] = x(0.015 + x) \approx 0.015x \approx 2.8 \times 10^{-13}$$

$$x = 1.9 \times 10^{-11} \text{ M}$$

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.

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

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



This is a buffer

$$\text{pH} = \text{p}K_a + \log \frac{[\text{CNO}^-]}{[\text{HCNO}]} = -\log(3.5 \times 10^{-4}) + \log \frac{15.2}{22.4} = 3.29$$