

Chemistry 112

- 3) If a 0.05 M solution of a weak acid produces a pH of 5.9, What is K_a for this weak acid?
- 7) Calculate the molar solubility of barium fluoride (BaF_2) in water.
 $K_{sp} = 1.6 \times 10^{-6}$.
Calculate the molar solubility of this compound in an aqueous solution that is 0.2 M NaF (sodium fluoride).
- 10) What is the pH of the following solutions?
a) 0.25 M HNO_3 b) 0.17 M $Ba(OH)_2$

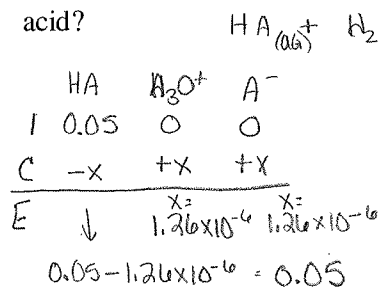
- 16) What is the concentration of Ca^{2+} ion remaining after CaCO_3 precipitates when 50.0 ml of 0.10 M CaCl_2 is added to 50.0 ml of 0.10 M Na_2CO_3 ? K_{sp} for CaCO_3 is 3.8×10^{-9} .

What is the pH of a 0.15 M solution of NH_3 ? $K_{\text{b}} = 1.78 \times 10^{-5}$

What is the PH of a 0.15 M solution of NH_4^+ ?

$$[H_3O^+] = 10^{-5.9} = 1.26 \times 10^{-6} = x$$

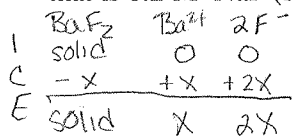
- 3) If a 0.05 M solution of a weak acid produces a pH of 5.9, What is K_a for this weak acid?



$$K_a = \frac{[1.26 \times 10^{-6}][1.26 \times 10^{-6}]}{[0.05]} = 3.2 \times 10^{-11}$$

- 7) Calculate the molar solubility of barium fluoride (BaF_2) in water. $K_{sp} = 1.6 \times 10^{-6}$. $BaF_2(s) \xrightleftharpoons{H_2O} Ba^{2+}_{(aq)} + 2F^-_{(aq)}$

Calculate the molar solubility of this compound in an aqueous solution that is 0.2 M NaF (sodium fluoride).



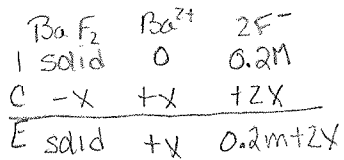
not on exam 2 - this is ch 17 - exam 3

$$K_{sp} = 1.6 \times 10^{-6} = [Ba^{2+}][F^-]^2 = [x][2x]^2 = 4x^3$$

$$\frac{1.6 \times 10^{-6}}{4} = \frac{4x^3}{4} \quad x^3 = 4 \times 10^{-7} \quad x = 7.4 \times 10^{-3} \text{ mol/L}$$

molar solubility = 7.4×10^{-3} mol/L

NaF complete dissociation



$$K_{sp} = [x][0.2 + 2x]^2$$

going to assume can ignore

$$= x[0.2]^2$$

$$1.6 \times 10^{-6} = 0.2x$$

$$x = 8.0 \times 10^{-6} = \text{molar solubility of } BaF_2 \text{ in } 0.2M \text{ NaF}$$

- 10) What is the pH of the following solutions?

a) 0.25 M HNO_3 strong acid

$$[H_3O^+] = 0.25M$$

$$pH = -\log(0.25)$$

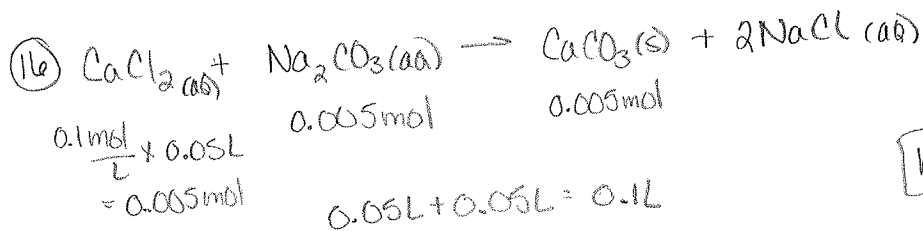
$$= 0.602$$

- b) 0.17 M $Ba(OH)_2$ strong base

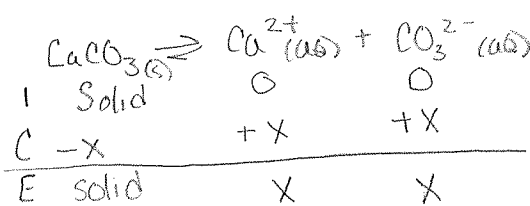
$$[OH^-] = 0.17M \times 2 = 0.34M$$

$$pOH = -\log(0.34) = 0.47$$

$$pH = 14 - 0.47 = 13.53$$



not on exam
ch 17



$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$

$$3.8 \times 10^{-9} = (x)(x) = x^2$$

$$x = 6.2 \times 10^{-5} \text{ mol/L}$$

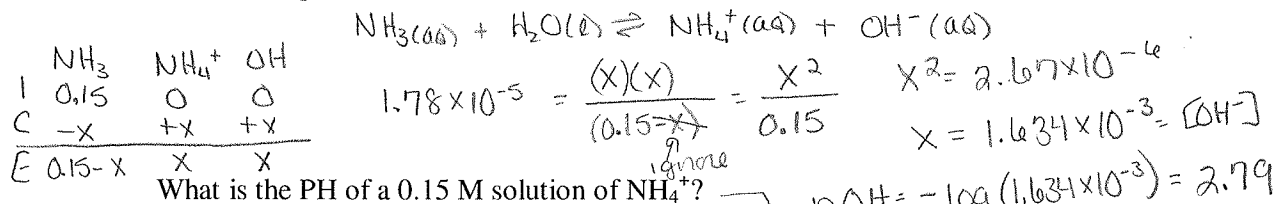
Dissolved $\text{Ca}^{2+} = 6.2 \times 10^{-5} \frac{\text{mol}}{\text{L}}$

- 16) What is the concentration of Ca^{2+} ion remaining after CaCO_3 precipitates when 50.0 ml of 0.10 M CaCl_2 is added to 50.0 ml of 0.10 M Na_2CO_3 ? K_{sp} for CaCO_3 is 3.8×10^{-9} .

not on exam 2

What is the pH of a 0.15 M solution of NH_3 ? $K_b = 1.78 \times 10^{-5}$

weak base

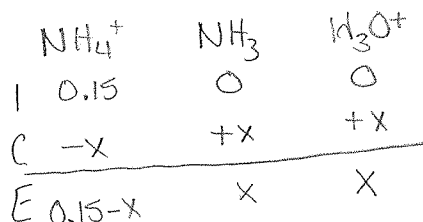


What is the PH of a 0.15 M solution of NH_4^+ ?



$\text{pOH} = -\log(1.634 \times 10^{-3}) = 2.79$
 $\text{pH} = 14 - 2.79 = \boxed{11.21}$

NH_4^+ is conjugate acid of NH_3 , so $K_a = \frac{1 \times 10^{-14}}{K_b} = \frac{1 \times 10^{-14}}{1.78 \times 10^{-5}} = 5.62 \times 10^{-10}$



$$5.62 \times 10^{-10} = \frac{[x][x]}{[0.15-x]} = \frac{x^2}{0.15}$$

ignore

$x^2 = 8.43 \times 10^{-11}$
 $x = 9.18 \times 10^{-6} = [\text{H}_3\text{O}^+]$
 $\text{pH} = -\log[9.18 \times 10^{-6}] = \boxed{5.04}$