

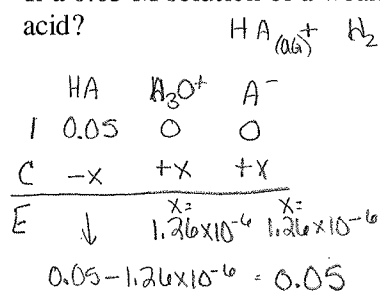


- 16) What is the concentration of  $\text{Ca}^{2+}$  ion remaining after  $\text{CaCO}_3$  precipitates when 50.0 ml of 0.10 M  $\text{CaCl}_2$  is added to 50.0 ml of 0.10 M  $\text{Na}_2\text{CO}_3$ ?  $K_{\text{sp}}$  for  $\text{CaCO}_3$  is  $3.8 \times 10^{-9}$ .

What is the pH of a 0.15 M solution of  $\text{NH}_3$ ?  $K_{\text{b}} = 1.78 \times 10^{-5}$

What is the PH of a 0.15 M solution of  $\text{NH}_4^+$ ?

- 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?



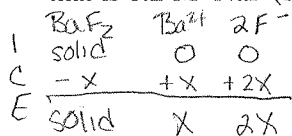
$[H_3O^+] = 10^{-5.9} = 1.26 \times 10^{-6} = x$

$$K_a = \frac{[A^-][H_3O^+]}{[HA]} = \frac{[1.26 \times 10^{-6}][1.26 \times 10^{-6}]}{[0.05]} = 3.2 \times 10^{-11}$$

- 7) Calculate the molar solubility of barium fluoride ( $BaF_2$ ) in water.  $K_{sp} = 1.6 \times 10^{-6}$ .

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

**Exam 3**



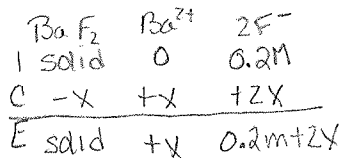
*not on exam 2 - this is ch 17 - exam 3*

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

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

*How much  $BaF_2$  dissolved*

NaF  
complete dissociation



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

*going to assume can ignore*

$$= x[0.2]^2$$

$$1.6 \times 10^{-6} = 0.2x \Rightarrow x = 8.0 \times 10^{-6} = \text{molar solubility of } BaF_2 \text{ in } 0.2M \text{ NaF}$$

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

a) 0.25 M  $HNO_3$  strong acid

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

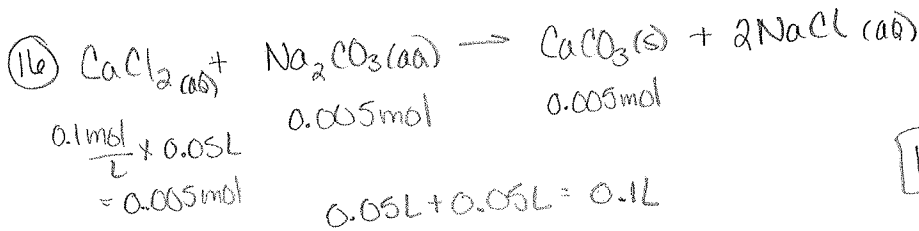
$$pH = -\log(0.25) = 0.602$$

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

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

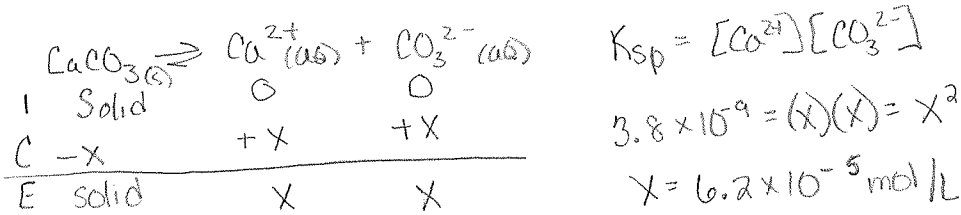
$$pOH = -\log(0.34) = 0.47$$

$$pH = 14 - 0.47 = 13.53$$



not on exam  
ch 17

Exam 3



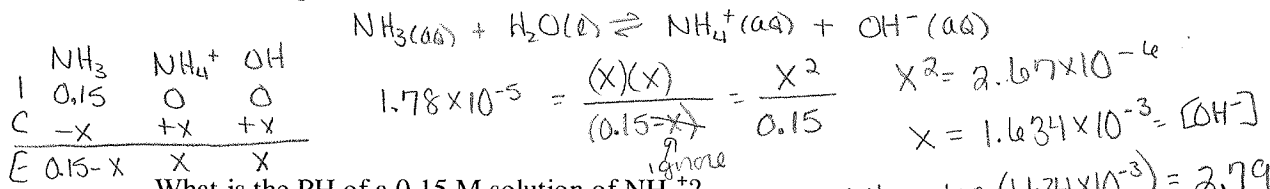
Dissolved  $\text{Ca}^{2+} = \frac{6.2 \times 10^{-5} \text{ mol}}{\text{L}}$

16) What is the concentration of  $\text{Ca}^{2+}$  ion remaining after  $\text{CaCO}_3$  precipitates when 50.0 ml of 0.10 M  $\text{CaCl}_2$  is added to 50.0 ml of 0.10 M  $\text{Na}_2\text{CO}_3$ ?  $K_{sp}$  for  $\text{CaCO}_3$  is  $3.8 \times 10^{-9}$ .

not on exam 2

What is the pH of a 0.15 M solution of  $\text{NH}_3$ ?  $K_b = 1.78 \times 10^{-5}$

weak base

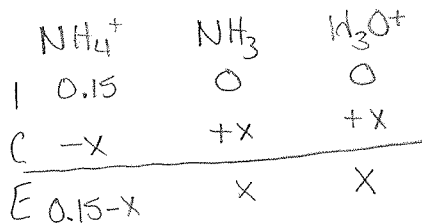


What is the PH of a 0.15 M solution of  $\text{NH}_4^+$ ?

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



$\text{NH}_4^+$  is conjugate acid of  $\text{NH}_3$ , so  $K_a = \frac{1 \times 10^{-14}}{K_b} = \frac{1 \times 10^{-14}}{1.78 \times 10^{-5}} = 5.62 \times 10^{-10}$



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

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