

#82 Notes **Unit 10: Acids & Bases**  
Ch. Acids, Bases, and Solubility  
Ch. Acids/ Bases

\*\* Acids: start with "H" or end in  $-\text{CO}_2\text{H}$  ( $-\text{COOH}$ ) and have a sour taste.

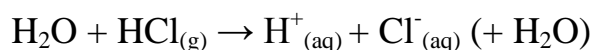
\*\* Bases: end in  $-\text{OH}$  (positive ion in front) or end in  $-\text{NH}$ ,  $-\text{NH}_2$ ,  $\text{NH}_3$

They have a bitter taste and feel slippery.

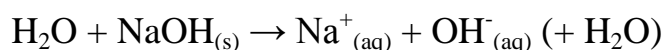
\* 3 Acid/Base Theories

I. Arrhenius Theory

a) Acids produce  $\text{H}^+$  in a water solution



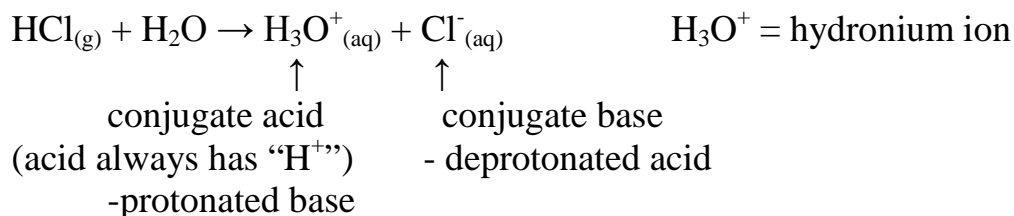
b) Bases produce  $\text{OH}^-$  in a water solution



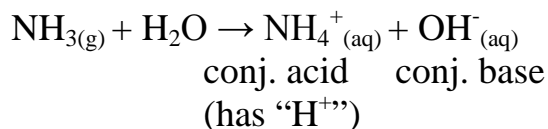
Acids, Bases and Salts are Electrolytes (Ion forming in water).

II. Brønsted-Lowry Theory

a) Acids donate protons ( $\text{H}^+$ ) ← lose  $1\text{e}^-$ , 1 p left, 0 neutrons



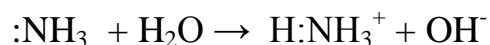
b) Bases accept protons



III. Lewis Theory

a) Acids are  $\text{e}^-$  pair acceptors.

b) Bases are  $\text{e}^-$  pair donors.



\* Bases have extra electrons.

\* Strong acids/bases completely ionize (fall apart).

Acids: HCl, HNO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>, HClO<sub>4</sub> (if #O - #H ≥ 2, then strong acid)

Bases: all Grp. 1 metals + OH<sup>-</sup> (like NaOH) plus Ca(OH)<sub>2</sub>, Ba(OH)<sub>2</sub>, Sr(OH)<sub>2</sub>  
- no K's

\* Weak acids/bases only partially ionize.

(Equilibriums – the larger the K, the more H<sub>3</sub>O<sup>+</sup>/OH<sup>-</sup> the stronger the acid/base.)

#### IV. pH scale

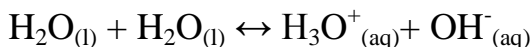
- is a measure of the acidity(H<sub>3</sub>O<sup>+</sup> or H<sup>+</sup>)

← 0 – 7 – 14 →  
acid base

$$* \text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$* \text{pOH} = -\log[\text{OH}^-]$$

$$** \text{H}_3\text{O}^+ = \text{H}^+$$



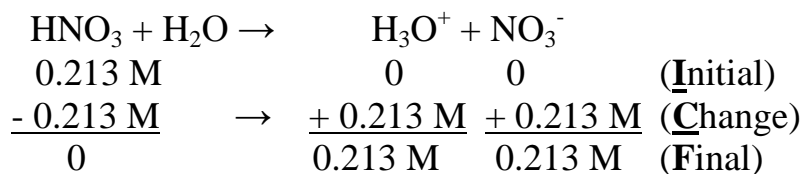
$$* \text{K}_w = 1 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

-log ( )

$$* 14 = \text{pH} + \text{pOH}$$

Ex. 1) Find [H<sub>3</sub>O<sup>+</sup>], [OH<sup>-</sup>], pH, & pOH for a **strong acid**:

a) 0.213 M HNO<sub>3</sub> (no K, so strong acid)



$$\text{H}_3\text{O}^+ = 0.213 \text{ M}$$

$$\text{pH} = -\log(\text{H}_3\text{O}^+) = -\log(0.213 \text{ M}) = -(-0.672) = \mathbf{0.672}$$

$$\text{K}_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$1 \times 10^{-14} = (0.213 \text{ M})(\text{OH}^-)$$

$$\mathbf{4.69 \times 10^{-14} \text{ M} = [\text{OH}^-]} \quad \text{pOH} = -\log(\text{OH}^-) = -\log(4.69 \times 10^{-14} \text{ M}) \\ = -(-13.329) = \mathbf{13.329}$$

\*\*The whole number is the power and is not significant!

OR

$$\begin{aligned} \text{From pH} = \mathbf{0.672} \quad \text{pH} + \text{pOH} &= 14 \\ 0.672 + \text{pOH} &= 14 \\ \text{pOH} &= \mathbf{13.328} \end{aligned}$$

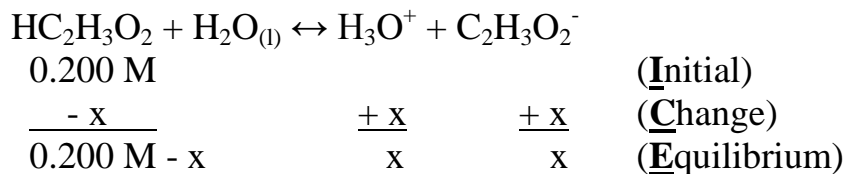
$$\begin{aligned} \text{pOH} &= -\log(\text{OH}^-) \\ 13.328 &= -\log(\text{OH}^-) \\ -13.328 &= \log(\text{OH}^-) \quad \text{move (-) before } 10^x \end{aligned}$$

$$\mathbf{4.70 \times 10^{-14} = [\text{OH}^-]}$$

To determine if acidic or basic, look at pH (pH < 7 then acidic)  
Or look at  $[\text{H}_3\text{O}^+]$  and  $[\text{OH}^-]$ :  
if  $[\text{H}_3\text{O}^+]$  is larger, it is acidic,  
if  $[\text{OH}^-]$  is larger, it is basic.



Ex. 1) a) What is the pH of a 0.200 M solution of acetic acid ( $\text{CH}_3\text{COOH} = \text{HC}_2\text{H}_3\text{O}_2$ )?  
 $K = 1.8 \times 10^{-5}$



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} \quad \text{no liquids: } \text{H}_2\text{O}_{(l)}$$

$$1.8 \times 10^{-5} = \frac{(x)(x)}{0.200 \text{ M} - x} \quad \text{If } K_a \text{ is smaller than the Concentration by } \mathbf{10^3 \text{ or more}}, \text{ assume } x \text{ is small. (} 0.200 - x \text{ would equal just } 0.200\text{)}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.200 \text{ M}}$$

$$\mathbf{1.9 \times 10^{-3} \text{ M} = x = [\text{H}_3\text{O}^+]}$$

$$\mathbf{\text{pH} = -\log(1.9 \times 10^{-3}) = 2.72}$$

check:

$$\frac{(1.9 \times 10^{-3})^2}{0.198} = 1.82 \times 10^{-5} \approx 1.8 \times 10^{-5} \quad \underline{\text{OK}}$$

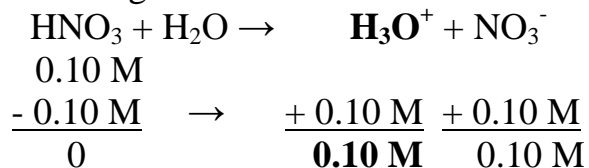
1b) Find the % dissociation.

$$\% \text{ dissociation} = \frac{\text{amount dissociated}}{\text{initial concentration}} \times 100 \quad (\text{amount dissociated} = x)$$

$$= \frac{1.9 \times 10^{-3}}{0.200 \text{ M}} \times 100 = \mathbf{0.95 \%}$$

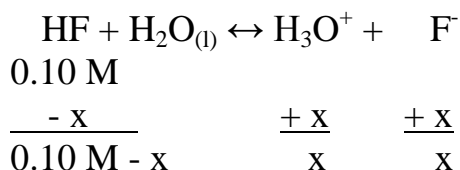
Ex. 2) Calculate the pH of a solution containing 0.10 M  $\text{HNO}_3$  and 0.10 M HF.  
**(strong & weak acids)**

$\text{HNO}_3$  is a strong acid.



$$\mathbf{[\text{H}_3\text{O}^+] = 0.10 \text{ M}}$$

HF is a weak acid.



$$K_a = 7.2 \times 10^{-4} = \frac{(x)^2}{(0.10 - x)} \quad x \text{ is small}$$

$$7.2 \times 10^{-4} = \frac{x^2}{0.10}$$

$$8.5 \times 10^{-3} \text{ M} = x = [\text{H}_3\text{O}^+]$$

So between the 2 acids  $[\text{H}_3\text{O}^+] = 0.10 \text{ M} + 8.5 \times 10^{-3} \text{ M} = \mathbf{0.11 \text{ M}}$  **pH = 0.96**

If HF was a little less, it can be ignored (just look at the strong acid).

Ex. 3) Calculate  $[\text{H}^+]$ , if 25 ml of 0.020 M HCl is added to 35 ml of 0.040 M HClO<sub>4</sub>.  
**(both are strong acids)**

HCl

M = mol/L

$$0.020 \text{ M} = \text{mol} / 0.025 \text{ L}$$

HClO<sub>4</sub>

$$0.040 \text{ M} = \text{mol} / 0.035 \text{ L}$$

$$0.00050 = \text{mol HCl} = \text{mol H}_3\text{O}^+ \quad 0.0014 = \text{mol HClO}_4 = \text{mol H}_3\text{O}^+$$

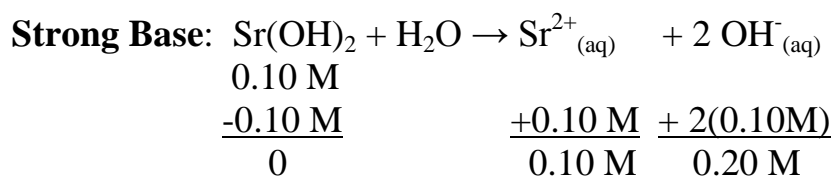
$$M = \text{mol} / L = \frac{0.00050 \text{ mol} + 0.0014 \text{ mol}}{0.025 \text{ L} + 0.035 \text{ L}} = \frac{0.0019 \text{ mol}}{0.060 \text{ L}} = \mathbf{0.032 \text{ M H}^+}$$

Then find pH, OH<sup>-</sup> etc.

\*If it asks for major species present, list concentrations of all ions/compounds.

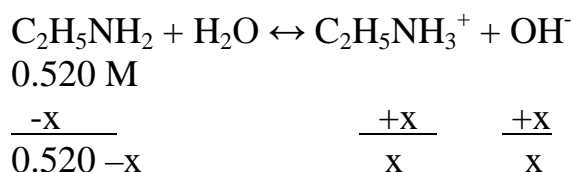
\*\* Don't forget, if it is a strong acid the concentration of acid = M H<sub>3</sub>O<sup>+</sup>

Ex. 1) Find pH of 0.10 M Sr(OH)<sub>2</sub>.



$$[\text{OH}^-] = 0.20 \text{ M} \quad \text{pOH} = 0.70 \quad \text{pH} = \mathbf{13.30}$$

Ex. 2) What is the pH of a 0.520 M solution of ethyl amine, C<sub>2</sub>H<sub>5</sub>NH<sub>2</sub>?  
(**weak base**) K = 5.6 X10<sup>-4</sup>



$$K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]}$$

$$5.6 \times 10^{-4} = \frac{(x^2)}{(0.520 - x)} \quad \text{x is small}$$

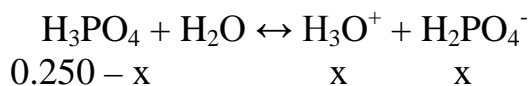
$$1.71 \times 10^{-2} = x = \text{OH}^-$$

$$\text{pOH} = 1.77 \quad 14 - 1.77 = \mathbf{12.23} = \text{pH}$$

### VII. Polyprotic Acids

-are acids that can furnish more than 1 H<sup>+</sup> (H<sub>2</sub>SO<sub>4</sub>, H<sub>3</sub>PO<sub>4</sub>, H<sub>3</sub>AsO<sub>4</sub> etc., see table in textbook)

Ex. 1) Calculate the pH and concentration of all species in a 0.250 M solution of phosphoric acid. K<sub>1</sub> = 7.5 X10<sup>-3</sup>, K<sub>2</sub> = 6.2 X10<sup>-8</sup>, K<sub>3</sub> = 4.8 X10<sup>-13</sup>



$$K_{a1} = 7.5 \times 10^{-3} = \frac{x^2}{0.250 - x} \quad \text{Quadratic}$$

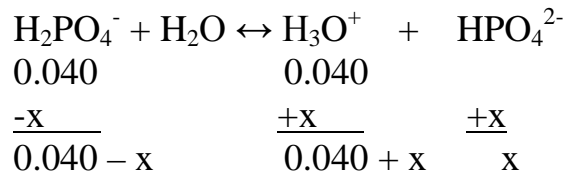
$$1.88 \times 10^{-3} - 7.5 \times 10^{-3}x - x^2 = 0$$

$$\frac{7.5 \times 10^{-3} \pm \sqrt{(-7.5 \times 10^{-3})^2 - 4(-1)(1.88 \times 10^{-3})}}{2(-1)} = \frac{7.5 \times 10^{-3} \pm 0.087}{-2} = 0.040$$

$$x = 0.040$$

$$\mathbf{H_3PO_4 = 0.250 - x = 0.21 \text{ M}}$$

$$\mathbf{H_2PO_4^- = H_3O^+ = x = 0.040 \text{ M}}$$



$$K_{a2} = 6.2 \times 10^{-8} = \frac{(0.040 + x)(x)}{(0.040 - x)}$$

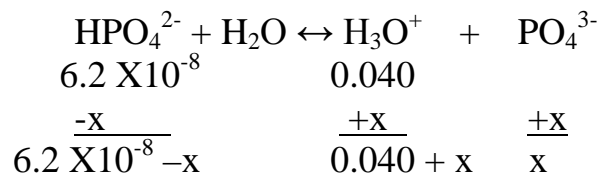
$$6.2 \times 10^{-8} = \frac{(0.040)(x)}{(0.040)}$$

$$6.2 \times 10^{-8} = x$$

$$\text{H}_2\text{PO}_4^- = 0.040 - x = 0.040 \text{ M}$$

$$\text{H}_3\text{O}^+ = 0.040 + x = 0.040 \text{ M}$$

$$\mathbf{HPO_4^{2-} = x = 6.2 \times 10^{-8} \text{ M}}$$



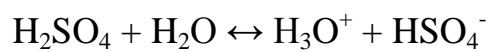
$$K_{a3} = 4.8 \times 10^{-13} = \frac{(0.040 + x)(x)}{(6.2 \times 10^{-8} - x)}$$

$$4.8 \times 10^{-13} = \frac{(0.040)(x)}{(6.2 \times 10^{-8})}$$

$$\mathbf{7.44 \times 10^{-19} = x = PO_4^{3-}}, \text{ x is too small to change } \text{H}_3\text{O}^+ \text{ or } \text{HPO}_4^{2-}$$

$$\text{H}_3\text{O}^+ = 0.040 \text{ M} \quad \mathbf{pH = 1.40}$$

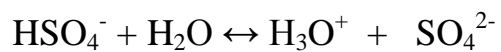
For 0.250 M H<sub>2</sub>SO<sub>4</sub>: K<sub>1</sub> = large, K<sub>2</sub> = 1.2 X10<sup>-2</sup>



0.250 M

$$\begin{array}{ccc} \frac{-0.250}{0} & \frac{+0.250}{0.250} & \frac{+0.250}{0.250} \end{array}$$

\*\*strong acid so [H<sub>3</sub>O<sup>+</sup>] = 0.250 M = [HSO<sub>4</sub><sup>-</sup>]



0.250 M            0.250 M

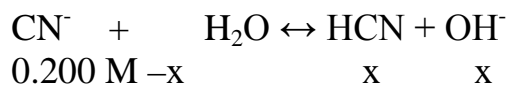
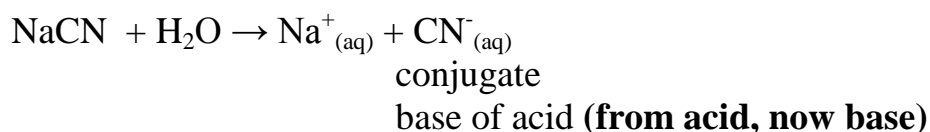
$$\begin{array}{ccc} \frac{-x}{0.250 - x} & \frac{+x}{0.250 + x} & \frac{+x}{x} \end{array}$$

then do K<sub>2</sub>





Ex. 2 ) Find pH of 0.200 M NaCN

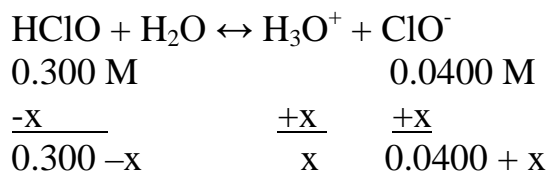


Find  $K_a$  of HCN ( $6.2 \times 10^{-10}$ ),  
use  $K_a \times K_b = K_w$  to get  $K_b$  ( $1.6 \times 10^{-5}$ )  
and solve.

Do you recognize these?  $\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl}$ , NaF,  $\text{H}_3\text{NOHCl}$ ,  $\text{Na}_2\text{SO}_4$ ,  $(\text{C}_6\text{H}_5)_2\text{NH}_2\text{Cl}$   
from base, from acid, from base, from acid, from base,  
now acid now base now acid now base now acid

### Ch. I. Common Ion Effect

Ex.1) What is the pH of a 0.300 M solution of HClO buffered with 0.0400 M NaClO?  
 $K$  of HClO =  $3.5 \times 10^{-8}$   
(buffered solutions resist changes in pH by having the acid & CB or the base & CA)

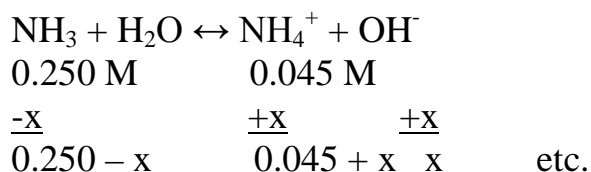


$$3.5 \times 10^{-8} = \frac{(x)(0.0400 + x)}{(0.300 - x)}$$

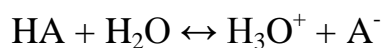
$$3.5 \times 10^{-8} = \frac{(x)(0.0400)}{(0.300)}$$

$$2.6 \times 10^{-7} = x = [\text{H}_3\text{O}^+] \quad \mathbf{pH = 6.58}$$

Ex. 2) 0.250 M  $\text{NH}_3$ , 0.045 M  $\text{NH}_4\text{Cl}$ ,  $K$  of  $\text{NH}_3 = 1.8 \times 10^{-5}$



## II. Henderson – Hasselbalch Equation



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

*log of both sides*

$$\log K_a = \log [\text{H}_3\text{O}^+] + \log ([\text{A}^-] / [\text{HA}])$$

$$-\log [\text{H}_3\text{O}^+] = -\log K_a + \log ([\text{A}^-] / [\text{HA}])$$

$$\text{pH} = \text{p}K_a + \log ([\text{A}^-] / [\text{HA}]) \quad \text{pH} = \text{p}K_a + \log ([\text{B}] / [\text{BH}^+])$$

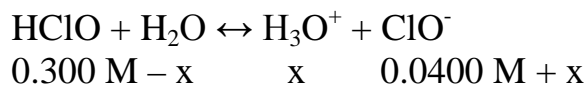
conjugate base  
acid

base  
conjugate acid

$$\uparrow$$

\*\*If base,  $K_w = K_a \times K_b$

Ex. 1) Notes Asst. #85 Common Ion ex. 1) 0.300 M HClO, 0.0400 M NaClO



$$\text{pH} = -\log (3.5 \times 10^{-8}) + \log ([0.0400 + x] / [0.300 - x]) \quad \text{x is small}$$

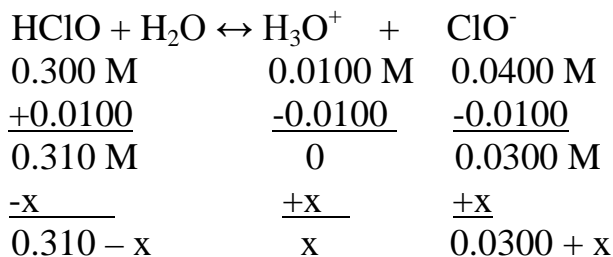
$K_a$

$$\text{pH} = -\log (3.5 \times 10^{-8}) + \log (0.0400 / 0.300) = 7.46 - 0.875 = \mathbf{6.58}$$

#86 Notes III. Buffered Solutions with Additives

A) Acid added to Acid Solutions

Ex. 1) Calculate the change in pH, if 0.0100 M HCl is added to 1 L of 0.300 M HClO and 0.0400 M NaClO.



\*\*If two concentrations on one side, subtract off the smallest amount.

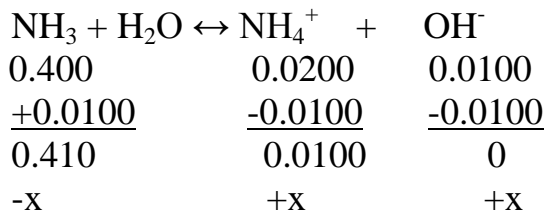
Then move over a "small" amount of x.

$$\text{pH} = \text{pK}_a + \log \left( \frac{[\text{ClO}^-]}{[\text{HClO}]} \right) \quad x \text{ is small}$$

$$\text{pH} = -\log(3.5 \times 10^{-8}) + \log(0.0300 / 0.310) = 7.46 - 1.014 = \mathbf{6.45}$$

B) Base added to Base Solutions

0.400 M NH<sub>3</sub>, 0.0200 M NH<sub>4</sub>Cl, 0.0100 M NaOH



K<sub>a</sub> X K<sub>b</sub> = K<sub>w</sub> to find K<sub>a</sub>, then use Henderson Hasselbalch

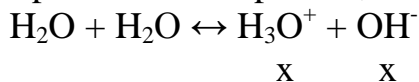
C) Acid or Base added to pure H<sub>2</sub>O

A strong acid or base totally dissociates. So concentration of acid or base = [H<sub>3</sub>O<sup>+</sup>] or [OH<sup>-</sup>], unless it has a very small concentration.

0.0100 M NaOH added to H<sub>2</sub>O = just 0.0100 M OH<sup>-</sup>, pOH = 2, pH = 12 (basic)

but if 1 X 10<sup>-11</sup> M NaOH, then

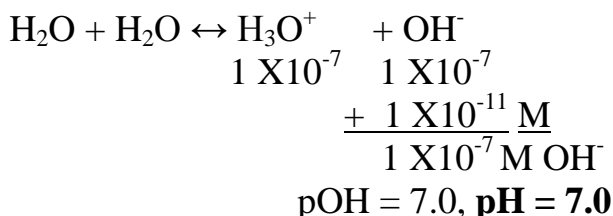
(If you use, 1 X 10<sup>-11</sup> M OH, then pOH = 11 and pH = 3, but OH<sup>-</sup> is not an acid!)



$$K_w = 1 \times 10^{-14} = [\text{H}_3\text{O}^+] [\text{OH}^-]$$

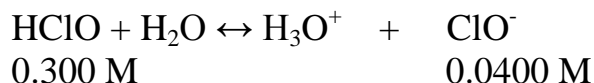
$$1 \times 10^{-14} = x^2$$

$$1 \times 10^{-7} = x = [\text{H}_3\text{O}^+] = [\text{OH}^-]$$

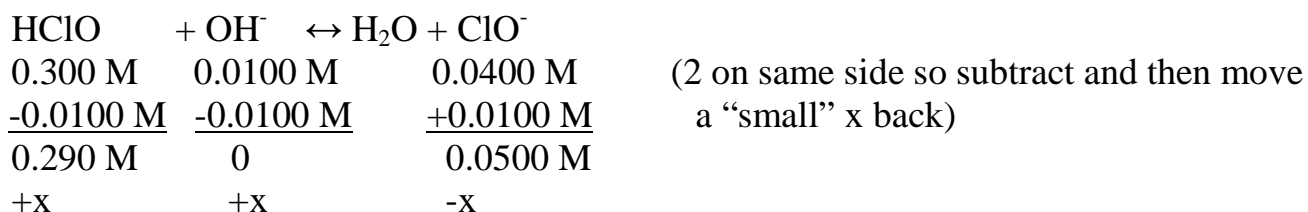


#### D) Strong Base added to Acid Solution

0.0100 M NaOH, 0.300 M HClO, 0.0400 M NaClO



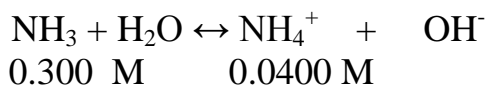
There is no place for the OH<sup>-</sup>, so must make new reaction.  
React Acid with strong Base!



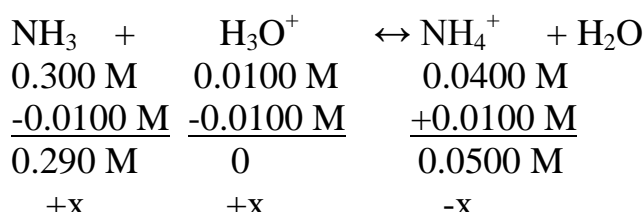
$$\begin{aligned} \text{pH} &= \text{pK}_a + \log \left( \frac{[\text{ClO}^-]}{[\text{HClO}]} \right) \\ \text{pH} &= -\log 3.5 \times 10^{-8} + \log (0.0500 / 0.290) = 7.46 - 0.763 = \mathbf{6.70} \end{aligned}$$

#### E) Strong Acid added to Base Solution

0.0100 M HCl, 0.300 M NH<sub>3</sub>, 0.0400 M NH<sub>4</sub>Cl



There is no place for H<sub>3</sub>O<sup>+</sup>.  
React Base with strong Acid!

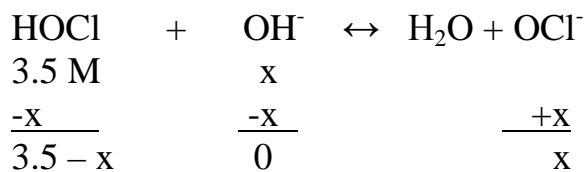


$$\begin{aligned} \text{pH} &= \text{pK}_a + \log \left( \frac{[\text{NH}_3]}{[\text{NH}_4^+]} \right) \\ \text{pH} &= -\log 5.6 \times 10^{-10} + \log (0.290 / 0.0500) \\ \text{pH} &= 9.25 + 0.76 = \mathbf{10.01} \end{aligned}$$

$$\begin{aligned} K_w &= K_a \times K_b \\ 1 \times 10^{-14} &= K_a (1.8 \times 10^{-5}) \\ 5.6 \times 10^{-10} &= K_a \end{aligned}$$

## #87 Notes IV. Working with Buffered Solutions

Ex. 1) What mass of solid NaOH must be added to 1 L of 3.5 M HOCl to produce a solution buffered at a pH of 5.00?  $K$  of HOCl =  $3.5 \times 10^{-8}$



$$\text{pH} = \text{pKa} + \log \left( \frac{[\text{OCl}^-]}{[\text{HOCl}]} \right)$$

$$5.00 = -\log 3.5 \times 10^{-8} + \log \left( \frac{x}{3.5 - x} \right)$$

$$5.00 = 7.46 + \log \left( \frac{x}{3.5 - x} \right)$$

$$10^x \{ -2.46 = \log \left( \frac{x}{3.5 - x} \right) \}$$

$$3.5 \times 10^{-3} = \left( \frac{x}{3.5 - x} \right)$$

$$0.012 - 3.5 \times 10^{-3}x = x$$

$$0.012 = 1.0035x$$

$$0.012 = x$$

**0.012 M OH<sup>-</sup>**

M = mol / L

0.012 M = mol / 1 L

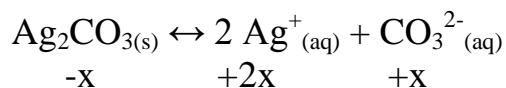
0.012 = mol OH<sup>-</sup>

$$\frac{0.012 \text{ mol NaOH}}{1 \text{ mol}} \left| \frac{39.997 \text{ g NaOH}}{1 \text{ mol}} \right. = \mathbf{0.48 \text{ g NaOH}}$$

## V. Solubility Equilibria

- describes the amount of solid that dissolves in a saturated solution.

Ex. 1) Calculate the solubility of Ag<sub>2</sub>CO<sub>3</sub> in mols per liter.  $K_{sp} = 8.1 \times 10^{-12}$



$$K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}]^1$$

$$K_{sp} = (2x)^2 (x)^1$$

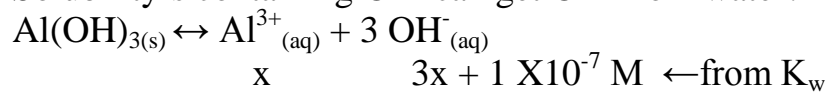
$$K_{sp} = 4x^3$$

$$8.1 \times 10^{-12} = 4x^3$$

$$1.3 \times 10^{-4} \text{ M} = x$$

**1.3 × 10<sup>-4</sup> mol/L**

Solubility's containing OH<sup>-</sup> can get OH<sup>-</sup> from water!



$$K_{sp} = [\text{Al}^{3+}]^1 [\text{OH}^{-}]^3$$

$$2 \times 10^{-32} = (x) (3x + 1 \times 10^{-7})^3$$

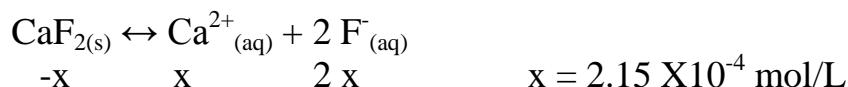
$$2 \times 10^{-32} = (x) (1 \times 10^{-7})^3 \quad x \text{ is small}$$

$$2 \times 10^{-11} \text{ M} = x$$

$$[\text{OH}^{-}] = 3x + 1 \times 10^{-7} = 6 \times 10^{-11} + 1 \times 10^{-7} = \mathbf{1 \times 10^{-7} \text{ M}}$$

\*\*If there is a small K, with OH<sup>-</sup>, then x is small and only the concentration of OH<sup>-</sup> in water matters.

Ex. 2) Calculate the  $K_{sp}$  of  $\text{CaF}_2$ , if its solubility is  $2.15 \times 10^{-4}$  mol/L.



$$K_{sp} = [\text{Ca}^{2+}]^1 [\text{F}^-]^2$$

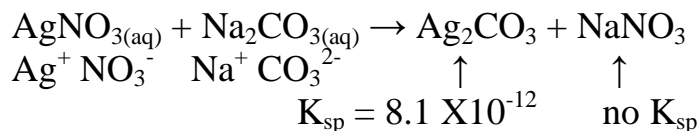
$$K_{sp} = (x)^1 (2x)^2$$

$$K_{sp} = 4x^3$$

$$K_{sp} = 4 (2.15 \times 10^{-4} \text{ mol/L})^3 = \mathbf{4.0 \times 10^{-11}}$$

## VI. Precipitation Reactions

Ex.1) What are the concentrations of all ions, if 25 ml of 0.100 M  $\text{AgNO}_3$  is added to 105 ml of 0.200 M  $\text{Na}_2\text{CO}_3$ ? Does a precipitate form?



a) Find concentration of ions in new volume:  $25 \text{ ml} + 105 \text{ ml} = 130 \text{ ml}$  (new solution)

$$ML = ML$$

$$(0.100 \text{ M AgNO}_3) (25 \text{ ml}) = M (130 \text{ ml})$$

$$\mathbf{0.0192 \text{ mol/L}} = M_{\text{AgNO}_3}$$

$$\frac{X1\text{Ag}^+}{0.0192 \text{ M Ag}^+} \quad \frac{X1\text{NO}_3^-}{0.0192 \text{ M NO}_3^-}$$

$$(0.200 \text{ M Na}_2\text{CO}_3) (105 \text{ ml}) = M (130 \text{ ml})$$

$$\mathbf{0.162 \text{ mol/L}} = M_{\text{Na}_2\text{CO}_3}$$

$$\frac{X2\text{Na}^+}{0.324 \text{ M Na}^+} \quad \frac{X1\text{CO}_3^{2-}}{0.162 \text{ M CO}_3^{2-}}$$

b) Put concentrations into  $K_{sp}$  equation:



$$K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}]^1 = (0.0192 \text{ M Ag}^+)^2 (0.162 \text{ M CO}_3^{2-})^1 = 5.97 \times 10^{-5}$$

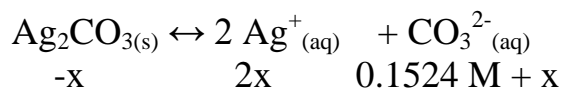
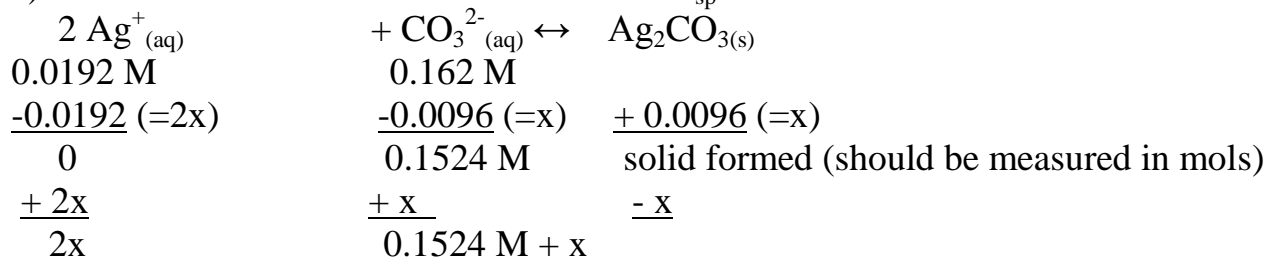
**$K_{sp}$  calculated =  $5.97 \times 10^{-5}$  >>> Real  $K_{sp}$  from book =  $8.1 \times 10^{-12}$**   
**for a saturated solution**

**since  $K_{sp}$  is more than saturated, solid forms (precipitate)**



The real  $K_{sp}$  determines how much of each ion is soluble in a saturated (or over saturated) solution.

c) Find how much solid forms and then use  $K_{sp}$ .



$$K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}]^1$$

$$8.1 \times 10^{-12} = (2x)^2 (0.1524 \text{ M} + x)^1 \quad x \text{ is small}$$

$$8.1 \times 10^{-12} = 4x^2 (0.1524)$$

$$3.64 \times 10^{-6} \text{ M} = x \quad 2x = [\text{Ag}^+] = 7.29 \times 10^{-6} \text{ M}$$

$$0.1524 \text{ M} + x = [\text{CO}_3^{2-}] = 0.152 \text{ M}$$

**The  $[\text{Na}^+]$  and  $[\text{NO}_3^-]$  would be the same as found above.**

### Ch. Free Energy and $K_p$

$$\Delta G = \Delta G^\circ + RT \ln Q_p$$

$\Delta G$  is at **non-equilibrium** conditions

$\Delta G^\circ$  is at equilibrium, calculated by

$$\Delta G_{\text{reaction}} = \sum n \Delta G_{\text{products}} - \sum n \Delta G_{\text{reactants}}$$

$$R = 8.31 \text{ J/mol}\cdot\text{K}$$

$Q_p$  is  $K_p$  at non-equilibrium conditions

$$\text{Remember: } K_p = K (RT)^{\Delta n}$$

**At equilibrium**  $\Delta G = 0$ , so  $0 = \Delta G^\circ + RT \ln K_p$

$$\Delta G^\circ = -RT \ln K_p$$

Ex. 1) Find  $\Delta G$  for  $\text{N}_{2(g)} + 3 \text{H}_{2(g)} \leftrightarrow 2 \text{NH}_{3(g)}$  at  $25^\circ\text{C}$ , if  $P_{\text{N}_2} = 0.050 \text{ atm}$ ,  $P_{\text{H}_2} = 0.0010 \text{ atm}$ , and  $P_{\text{NH}_3} = 0.012 \text{ atm}$ .

$$\Delta G^\circ_{\text{reaction}} = \sum n\Delta G_{\text{products}} - \sum n\Delta G_{\text{reactants}} \quad (\text{old } \Delta G \text{ energy table})$$

$$\Delta G^\circ = [ (2 \text{ mol NH}_3) ( -16.45 \text{ kJ/mol} ) ] -$$

$$[ (1 \text{ mol N}_2) (0 \text{ kJ/mol}) + (3 \text{ mol H}_2) (0 \text{ kJ/mol}) ]$$

$$\Delta G^\circ = -32.9 \text{ kJ}$$

$$Q_p = \frac{P_{\text{NH}_3}^2}{P_{\text{N}_2}^1 P_{\text{H}_2}^3} = \frac{(0.012 \text{ atm})^2}{(0.050 \text{ atm})^1 (0.0010 \text{ atm})^3} = 2.88 \times 10^6$$

$$\Delta G = \Delta G^\circ + RT \ln Q_p$$

$$\Delta G = \frac{(-32.9 \text{ kJ} \mid 1 \times 10^3 \text{ J})}{1 \text{ kJ}} + (8.31 \text{ J/mol}\cdot\text{K}) (298 \text{ K}) \ln (2.88 \times 10^6)$$

(14.873)

$$\Delta G = -32900 \text{ J} + 36832 \text{ J} = 3932 = \mathbf{3.9 \times 10^3 \text{ J or } 3.9 \text{ kJ}}$$

**\*End of Notes\*** (Assignments #89-90 are Review Assignments. There are no notes for these assignments.)