

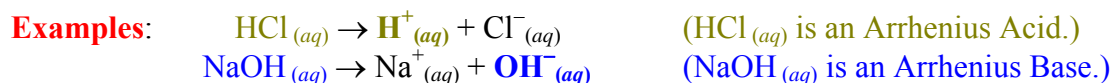
Unit 6: ACIDS-BASES AND SOLUBILITY EQUILIBRIA**Chapter 15: Acids and Bases****15.1: Brønsted Acids and Bases****Physical and Chemical Properties of Acid and Base**

Acids	Bases
Taste Sour (Citric Acids).	Taste Bitter.
Burning Sensation (Stomach Acid).	Feels Slippery (Detergent, Degreaser).
Corrosive with Metals (reacts to give off $H_2(g)$).	Alkaline in Nature (NaOH, Baking Soda).
Red litmus remains Red; Blue litmus turns Red.	Red litmus turns Blue; Blue litmus remains Blue.
Bromothymol Blue turns Yellow	Bromothymol Blue turns Blue.
Phenolphthalein turns Colourless.	Phenolphthalein turns Pink.
pH < 7	pH > 7



Conceptual Definition: - an explanation that attempts to describe why things are the way they are.

Arrhenius Concept: - acids are H^+ (proton) producers and bases are OH^- producers.

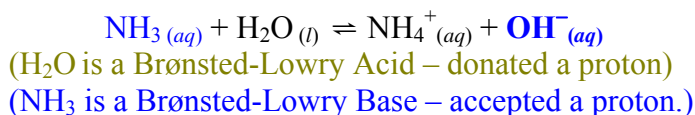
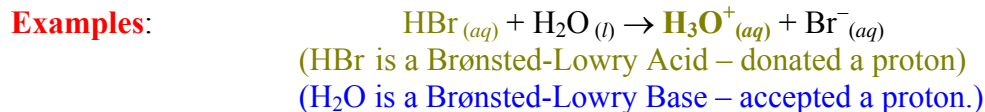


Brønsted-Lowry Model: - acids and bases **react with water** to **dissociate** where **acids are H^+ (proton) donors** and **bases are H^+ (proton) acceptors**.

- first proposed by Johannes Brønsted and Thomas Lowry.

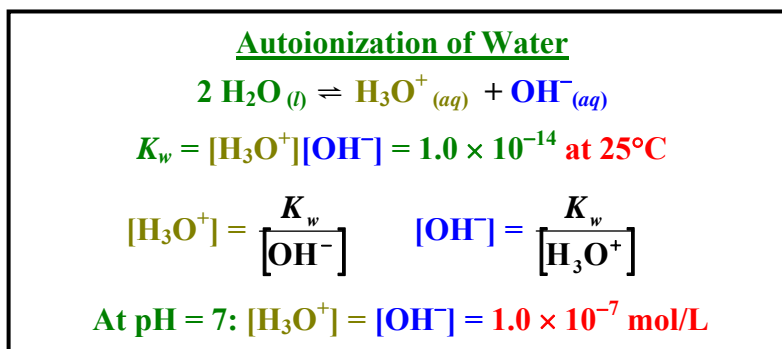
Hydronium Ion: - an ion formed when an **acid "donated" H^+ ion combined with a H_2O molecule** to form a **H_3O^+ ion (hydronium ion)**.

- essentially has the same function as a H^+ ion, but H_3O^+ denotes that we are using the Brønsted-Lowry model.



15.2: The Acid-Base Properties of Water

Autoionization of Water: - the process where water dissociates into hydronium and hydroxide ions.
 - water's dissociation constant is called **ion-product constant** ($K_w = 1.0 \times 10^{-14}$)
 - when the solution is **neutral (pH = 7)**, $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol/L}$



Example 1: At 25°C , $K_w = 1.0 \times 10^{-14}$.

- a. Using the ICE box, show that $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol/L}$ for a neutral solution.
 b. At 100°C , $K_w = 8.19 \times 10^{-13}$. What is the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ for a neutral solution at 100°C ?

a.

	$2 \text{H}_2\text{O} (l)$	\rightleftharpoons	$\text{H}_3\text{O}^+ (aq)$	$+$	$\text{OH}^- (aq)$
	$\text{H}_2\text{O} (l)$		$[\text{H}_3\text{O}^+]$		$[\text{OH}^-]$
Initial	----		0		0
Change	----		+x		+x
Equilibrium	-----		x		x

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

$$1.0 \times 10^{-14} = (x)(x)$$

$$x^2 = \sqrt{1.0 \times 10^{-14}}$$

$$x = 1.0 \times 10^{-7}$$

$x = [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$

b. At 100°C , using $K_w = 8.19 \times 10^{-13}$

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

$$8.19 \times 10^{-13} = (x)(x)$$

$$x^2 = \sqrt{8.19 \times 10^{-13}}$$

$$x = 2.86 \times 10^{-7}$$

$x = [\text{H}_3\text{O}^+] = [\text{OH}^-] = 2.86 \times 10^{-7} \text{ M}$

Example 2: Determine the $[\text{H}_3\text{O}^+]$ and/or $[\text{OH}^-]$ concentrations of the following solutions at 25°C .

a. $[\text{OH}^-] = 1.0 \times 10^{-4} \text{ mol/L}$

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

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}}$$

$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-10} \text{ mol/L}$

b. $[\text{H}_3\text{O}^+] = 5.0 \times 10^{-5} \text{ M}$

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

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-5}}$$

$[\text{OH}^-] = 2.0 \times 10^{-10} \text{ M}$

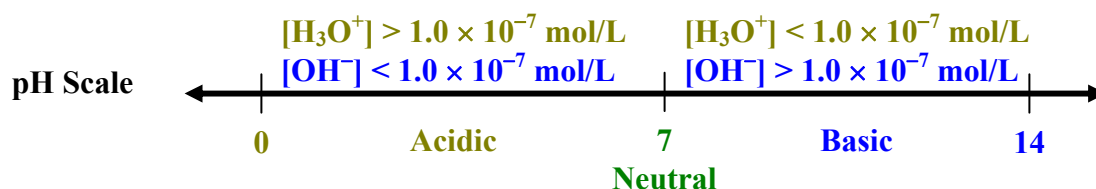
Assignment

15.1 pg. 686–687 #1 to 8
15.2 pg. 687 #9 to 11

15.3: pH – A Measure of Acidity

pH Scale: - a logarithmic scale to **measure the acidity (relative $[\text{H}_3\text{O}^+]$)** of a solution.

- the **lower the pH, the more acidic** (less basic) is the solution (**more $[\text{H}_3\text{O}^+]$ and less $[\text{OH}^-]$**).
- the **higher the pH, the more basic** (less acidic) is the solution (**less $[\text{H}_3\text{O}^+]$ and more $[\text{OH}^-]$**).
- **acidity is NOT the same as the acid strength**. Just because a solution has a low pH, it does not mean that it is a strong acid. (**Highly Acidic \neq Strong Acid**)
- it is normally reported between 0 to 14 (**with 7 as neutral**), but it **can be above 14 (very basic) or below 0 (very acidic)**.
- **an increase of 1 on a pH scale means a decrease of $[\text{H}_3\text{O}^+]$ by a factor of 10; an increase of 2 on a pH scale means a decrease of $[\text{H}_3\text{O}^+]$ by a factor of 100.**



Example: pH of Some Common Substances

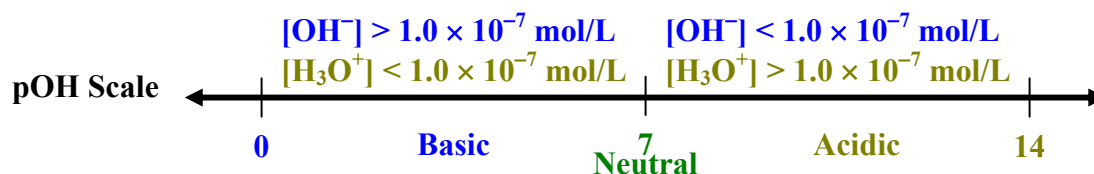
Substance	pH
1 M of HCl	0.00
Stomach Acid	2.00
Lemon Juice	2.50
Vinegar	3.00

Substance	pH
Milk	6.30
Rain Water	6.70
Pure Water	7.00
Blood	7.50

Substance	pH
1 M of Baking Soda (NaHCO_3)	9.68
Ammonia as Household Cleaner	12.00
1 M of NaOH	14.00

pOH Scale: - a logarithmic scale to **measure the basicity (relative $[\text{OH}^-]$)** of a solution.

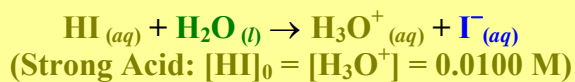
- the **lower the pOH, the more basic** (less acidic) is the solution (**more $[\text{OH}^-]$ and less $[\text{H}_3\text{O}^+]$**).
- the **higher the pOH, the less basic** (more acidic) is the solution (**less $[\text{OH}^-]$ and more $[\text{H}_3\text{O}^+]$**).
- **basicity is NOT the same as the base strength**. Just because a solution has a low pOH, it does not mean that it is a strong base. (**Highly Basic \neq Strong Base**)
- it is normally reported between 0 to 14 (**with 7 as neutral**), but it **can be above 14 (very acidic) or below 0 (very basic)**.
- **an increase of 1 on a pOH scale means a decrease of $[\text{OH}^-]$ by a factor of 10; an increase of 2 on a pOH scale means a decrease of $[\text{OH}^-]$ by a factor of 100.**



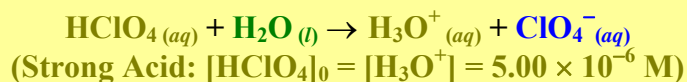
pH and pOH Scales	
$\text{pH} = -\log [\text{H}_3\text{O}^+]$	$\text{pOH} = -\log [\text{OH}^-]$
$\text{pH} + \text{pOH} = 14.00$	

Example 1: Calculate the pH and the pOH for the following solutions.

- a. 0.0100 mol/L of HI_(aq) (completely dissociates) b. 5.00×10^{-6} M of HClO_{4(aq)} (completely dissociates)



$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] & \text{pOH} &= 14 - \text{pH} \\ \text{pH} &= -\log(0.0100) & \text{pOH} &= 14 - 2.00 \\ \text{pH} &= 2.00 & \text{pOH} &= 12.00 \end{aligned}$$



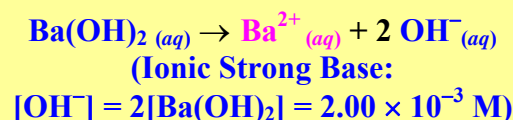
$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] & \text{pOH} &= 14 - \text{pH} \\ \text{pH} &= -\log(5.00 \times 10^{-6}) & \text{pOH} &= 14 - 5.30 \\ \text{pH} &= 5.30 & \text{pOH} &= 8.70 \end{aligned}$$

- c. 1.50 M of KOH_(aq)



$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] & \text{pH} &= 14 - \text{pOH} \\ \text{pOH} &= -\log(1.50) & \text{pH} &= 14 - (-0.176) \\ \text{pOH} &= -0.176 & \text{pH} &= 14.176 \end{aligned}$$

- d. 1.00×10^{-3} mol/L of Ba(OH)_{2(aq)}



$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] & \text{pH} &= 14 - \text{pOH} \\ \text{pOH} &= -\log(2.00 \times 10^{-3}) & \text{pH} &= 14 - (2.70) \\ \text{pOH} &= 2.70 & \text{pH} &= 11.30 \end{aligned}$$

Example 2: Calculate the $[\text{H}_3\text{O}^+]$ and the $[\text{OH}^-]$ for the following solutions.

- a. pH = 4.00

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] \\ [\text{H}_3\text{O}^+] &= 10^{-\text{pH}} = 10^{-4.00} \\ [\text{H}_3\text{O}^+] &= 1.00 \times 10^{-4} \text{ M} \end{aligned}$$

$$\begin{aligned} \text{pOH} &= 14 - \text{pH} \\ \text{pOH} &= 14 - 4.00 = 10.00 \end{aligned}$$

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ [\text{OH}^-] &= 10^{-\text{pOH}} = 10^{-10.00} \\ [\text{OH}^-] &= 1.00 \times 10^{-10} \text{ M} \end{aligned}$$

- b. pOH = 3.00

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ [\text{OH}^-] &= 10^{-\text{pOH}} = 10^{-3.00} \\ [\text{OH}^-] &= 1.00 \times 10^{-3} \text{ M} \end{aligned}$$

$$\begin{aligned} \text{pH} &= 14 - \text{pOH} \\ \text{pH} &= 14 - 3 = 11.00 \end{aligned}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] \\ [\text{H}_3\text{O}^+] &= 10^{-\text{pH}} = 10^{-11.00} \\ [\text{H}_3\text{O}^+] &= 1.00 \times 10^{-11} \text{ M} \end{aligned}$$

- c. pH = 12.83

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+] \\ [\text{H}_3\text{O}^+] &= 10^{-\text{pH}} = 10^{-12.83} \\ [\text{H}_3\text{O}^+] &= 1.48 \times 10^{-13} \text{ M} \end{aligned}$$

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

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{1.48 \times 10^{-13}}$$

$$[\text{OH}^-] = 0.0676 \text{ M}$$

- d. pOH = 9.67

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ [\text{OH}^-] &= 10^{-\text{pOH}} = 10^{-9.67} \\ [\text{OH}^-] &= 2.14 \times 10^{-10} \text{ M} \end{aligned}$$

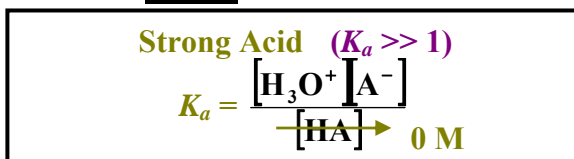
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{2.14 \times 10^{-10}}$$

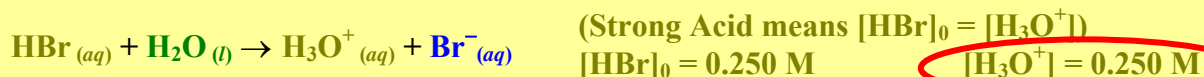
$$[\text{H}_3\text{O}^+] = 4.68 \times 10^{-5} \text{ mol/L}$$

15.4: Strength of Acids and Bases**Strong Acids:** - acids that **dissociate completely (100%) in water.**

- when the **acid dissociation constant (K_a) is much greater than 1.** ($K_a \gg 1$)
- **the equilibrium position strongly favours the products.**
- at equilibrium, the original acid concentration, $[HA]_{eq} \approx 0$; $[H_3O^+]_{eq} = [A^-]_{eq} = [HA]_0$.
- the **conjugate base, A^-** , of a strong acid, HA **is itself a weak base** (cannot easily accept protons to do the reverse reaction).
- **Note: Strong Acids DO NOT MEAN that they are VERY CORROSIVE. It is the $[H_3O^+]$ that defines acidity.**



	[HA]		[H ₃ O ⁺]	[A ⁻]
Initial	x		0	0
Change	$-x$		$+x$	$+x$
Equilibrium	0		x	x

Examples: Strong Acids: HClO₄ (aq), HI (aq), HBr (aq), HCl (aq), H₂SO₄ (aq) and HNO₃ (aq)**Example 2:** Write the dissociation reaction of 0.250 M of HBr (aq) and determine its [H₃O⁺].**Weak Acids:** - acids that **dissociate LESS than 100% in water.**

- when the **acid dissociation constant (K_a) is less than 1.** ($K_a < 1$)
- **the equilibrium position strongly favours the reactants.**
- at equilibrium, the hydronium concentration is much less than the original acid concentration, $[HA]_{eq} > [H_3O^+]_{eq}$ or $[HA]_0 \approx [HA]_{eq}$.
- the **conjugate base, A^-** , of a weak acid **is itself a stronger weak base** (can easily accept protons to do the reverse reaction).
- **Note: Weak Acids DO NOT MEAN that they are NOT CORROSIVE. It is the $[H_3O^+]$ that defines acidity. At a high enough concentration, a weak acid can be corrosive.**



	[HA]		[H ₃ O ⁺]	[A ⁻]
Initial	x		0	0
Change	$-y$ (where $y \ll x$)		$+y$	$+y$
Equilibrium	$(x - y) \approx x$		y	y

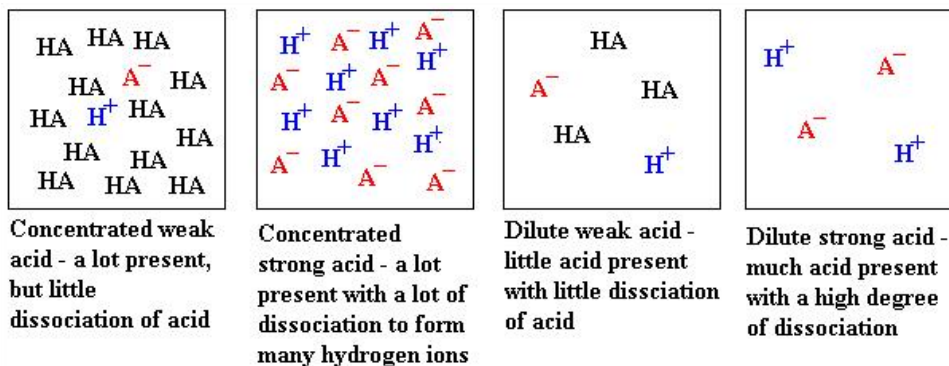
Examples: Some Weak Acids: HOCCOOH (aq), H₂SO₃ (aq), HSO₄⁻ (aq), H₃PO₄ (aq), HNO₂ (aq), H₃C₆H₅O₇ (aq), HF (aq), HCOOH (aq), C₆H₈O₆ (aq), C₆H₅COOH (aq), CH₃COOH (aq), H₂CO₃ (aq), H₂S (aq), HOCl (aq), HCN (aq), NH₄⁺ (aq), and H₃BO₃ (aq)

K_a and Relative Strength of Some Common Acids and Bases at 25°C

Acid Name	Acid Formula	Conjugate Base Formula	K_a
perchloric acid	$\text{HClO}_4 (aq)$	$\text{ClO}_4^- (aq)$	Very Large
hydroiodic acid	$\text{HI} (aq)$	$\text{I}^- (aq)$	Very Large
hydrobromic acid	$\text{HBr} (aq)$	$\text{Br}^- (aq)$	Very Large
hydrochloric acid	$\text{HCl} (aq)$	$\text{Cl}^- (aq)$	Very Large
sulfuric acid	$\text{H}_2\text{SO}_4 (aq)$	$\text{HSO}_4^- (aq)$	Very Large
nitric acid	$\text{HNO}_3 (aq)$	$\text{NO}_3^- (aq)$	Very Large
hydronium ion	$\text{H}_3\text{O}^+ (aq)$	$\text{H}_2\text{O} (l)$	1
oxalic acid	$\text{HOOC}(\text{COOH}) (aq)$	$\text{HOOC}(\text{COO})^- (aq)$	6.5×10^{-2}
sulfurous acid	$\text{H}_2\text{SO}_3 (aq)$	HSO_3^-	1.5×10^{-2}
hydrogen sulfate ion	$\text{HSO}_4^- (aq)$	SO_4^{2-}	1.2×10^{-2}
chlorous acid	$\text{HClO}_2 (aq)$	$\text{ClO}_2^- (aq)$	1.2×10^{-2}
phosphoric acid	$\text{H}_3\text{PO}_4 (aq)$	$\text{H}_2\text{PO}_4^- (aq)$	7.5×10^{-3}
arsenic acid	$\text{H}_3\text{AsO}_4 (aq)$	$\text{H}_2\text{AsO}_4^- (aq)$	5×10^{-3}
monochloroacetic acid	$\text{HC}_2\text{H}_2\text{ClO}_2$	$\text{C}_2\text{H}_2\text{ClO}_2^- (aq)$	1.35×10^{-3}
citric acid	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7 (aq)$	$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^- (aq)$	8.4×10^{-4}
hydrofluoric acid	$\text{HF} (aq)$	$\text{F}^- (aq)$	7.2×10^{-4}
nitrous acid	$\text{HNO}_2 (aq)$	$\text{NO}_2^- (aq)$	4.0×10^{-4}
methanoic (formic) acid	$\text{HCOOH} (aq)$	$\text{HCOO}^- (aq)$	1.8×10^{-4}
lactic acid	$\text{HC}_3\text{H}_5\text{O}_3 (aq)$	$\text{C}_3\text{H}_5\text{O}_3^- (aq)$	1.38×10^{-4}
ascorbic acid (vitamin C)	$\text{H}_2\text{C}_6\text{H}_6\text{O}_6 (aq)$	$\text{HC}_6\text{H}_6\text{O}_6^- (aq)$	7.9×10^{-5}
benzoic acid	$\text{C}_6\text{H}_5\text{COOH} (aq)$	$\text{C}_6\text{H}_5\text{COO}^- (aq)$	6.4×10^{-5}
hydrogen oxalate ion	$\text{HOOC}(\text{COO})^- (aq)$	$\text{OOC}(\text{COO})^{2-} (aq)$	6.1×10^{-5}
ethanoic (acetic) acid	$\text{CH}_3\text{COOH} (aq)$	$\text{CH}_3\text{COO}^- (aq)$	1.8×10^{-5}
dihydrogen citrate ion	$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^- (aq)$	$\text{HC}_6\text{H}_5\text{O}_7^{2-} (aq)$	1.8×10^{-5}
hydrated aluminum (III) ion	$[\text{Al}(\text{H}_2\text{O})_6]^{3+}$	$[\text{AlOH}(\text{H}_2\text{O})_5]^{2+}$	1.4×10^{-5}
propanoic acid	$\text{C}_2\text{H}_5\text{COOH} (aq)$	$\text{C}_2\text{H}_5\text{COO}^- (aq)$	1.3×10^{-5}
hydrogen citrate ion	$\text{HC}_6\text{H}_5\text{O}_7^{2-} (aq)$	$\text{C}_6\text{H}_5\text{O}_7^{3-} (aq)$	4.0×10^{-6}
carbonic acid	$\text{H}_2\text{CO}_3 (aq)$	$\text{HCO}_3^- (aq)$	4.3×10^{-7}
hydrosulfuric acid	$\text{H}_2\text{S} (aq)$	$\text{HS}^- (aq)$	1.0×10^{-7}
hydrogen sulfite ion	$\text{HSO}_3^- (aq)$	$\text{SO}_3^{2-} (aq)$	1.0×10^{-7}
dihydrogen arsenate ion	$\text{H}_2\text{AsO}_4^- (aq)$	$\text{HAsO}_4^{2-} (aq)$	8×10^{-8}
dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^- (aq)$	$\text{HPO}_4^{2-} (aq)$	6.2×10^{-8}
hypochlorous acid	$\text{HOCl} (aq)$	$\text{OCl}^- (aq)$	3.5×10^{-8}
hypobromous acid	$\text{HOBr} (aq)$	$\text{OBr}^- (aq)$	2×10^{-9}
hydrocyanic acid	$\text{HCN} (aq)$	$\text{CN}^- (aq)$	6.2×10^{-10}
hydrogen arsenate ion	$\text{HAsO}_4^{2-} (aq)$	$\text{AsO}_4^{3-} (aq)$	6×10^{-10}
boric acid	$\text{H}_3\text{BO}_3 (aq)$	$\text{H}_2\text{BO}_3^- (aq)$	5.8×10^{-10}
ammonium ion	$\text{NH}_4^+ (aq)$	$\text{NH}_3 (aq)$	5.6×10^{-10}
Phenol	$\text{C}_6\text{H}_5\text{OH} (aq)$	$\text{C}_6\text{H}_5\text{O}^- (aq)$	1.6×10^{-10}
hydrogen carbonate ion	$\text{HCO}_3^- (aq)$	$\text{CO}_3^{2-} (aq)$	5.6×10^{-11}
hypoiodous acid	$\text{HOI} (aq)$	$\text{OI}^- (aq)$	2×10^{-11}
hydrogen ascorbate ion	$\text{HC}_6\text{H}_6\text{O}_6^- (aq)$	$\text{C}_6\text{H}_6\text{O}_6^{2-} (aq)$	1.6×10^{-12}
hydrogen phosphate ion	$\text{HPO}_4^{2-} (aq)$	$\text{PO}_4^{3-} (aq)$	4.8×10^{-13}
water (55.49 mol/L)	$\text{H}_2\text{O} (l)$	$\text{OH}^- (aq)$	$1.0 \times 10^{-14} = K_w$
hydrogen sulfide ion	$\text{HS}^- (aq)$	$\text{S}^{2-} (aq)$	$\sim 10^{-19}$

Increasing Acid Strength

Increasing Base Strength

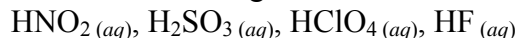


Check out Strong and Weak Acids Animations at
<http://www.sgc.peachnet.edu/users/larnold/WWW/courses/1212/rev1212.html>

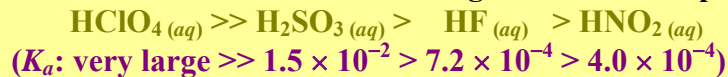
Relative Strength of Acids and Conjugate Bases:

1. The **stronger the acid** (the bigger the value of K_a), the **weaker its conjugate base**.
2. The **weaker the acid** (the smaller the value of K_a), the **stronger its conjugate base**.

Example 2: Order the following acids from the strongest to the weakest.



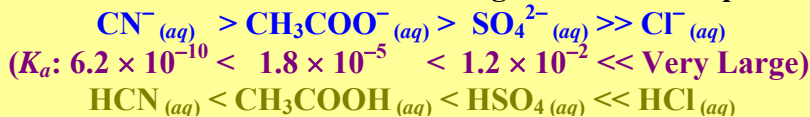
According to the Acid and Bases Relative Strength Table on the previous page:



Example 3: Order the following conjugate base from the strongest to the weakest.



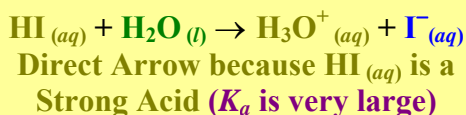
According to the Acids and Bases Relative Strength Table on the previous page:



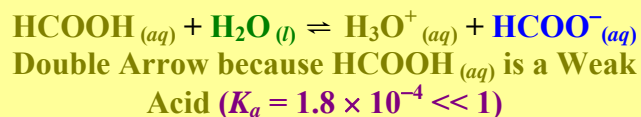
Monoprotic Acids: - acids that can donate a maximum of one proton.

Example 3: Write the dissociation reaction for the following monoprotic acids.

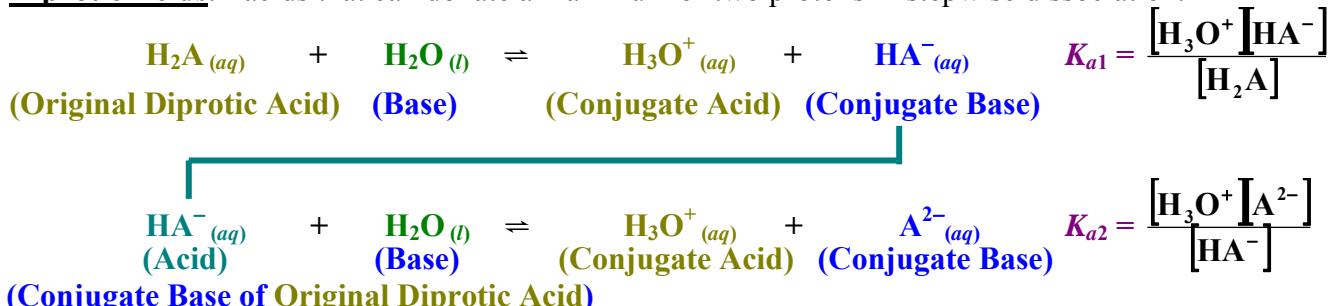
a. $\text{HI}(aq)$



b. $\text{HCOOH}(aq)$

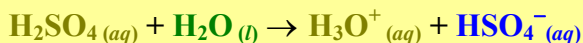


Diprotic Acids: - acids that can donate a maximum of two protons in stepwise dissociation.

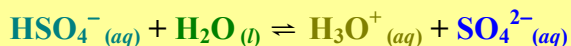


Example 4: Write the stepwise dissociation reaction for the following diprotic acids.

a. $\text{H}_2\text{SO}_4(aq)$

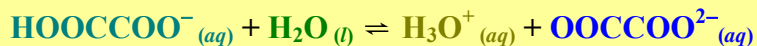
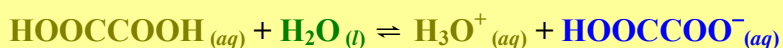


Direct Arrow because $\text{H}_2\text{SO}_4(aq)$ is a Strong Acid (K_{a1} is very large)



Double Arrow because $\text{HSO}_4^-(aq)$ is a Weak Acid ($K_{a2} = 1.2 \times 10^{-2} \ll 1$)

b. $\text{HOOC}\text{COOH}(aq)$



Double Arrow because both $\text{HOOC}\text{COOH}(aq)$ and HOOCCOO^- are Weak Acids (K_{a1} and $K_{a2} \ll 1$)

Amphoteric Substances: - chemical species that can be an acid or a base.

- all intermediate species of a diprotic acid is an amphoteric substance.

Examples: Some Amphoteric Substances: $\text{HOOC}\text{COO}^-(aq)$, $\text{HSO}_4^-(aq)$, $\text{HSO}_3^-(aq)$, $\text{HCO}_3^-(aq)$, $\text{HS}^-(aq)$, $\text{HC}_6\text{H}_6\text{O}_6^-(aq)$, and $\text{H}_2\text{O}(l)$

Calculating the pH of Strong Acid Solutions

Major Species: - the predominant species of an acid or a base after dissociation.

- in acid and base dissociation, because they are aqueous, water ($\text{H}_2\text{O}(l)$) is always listed as a major species.

Major Species of a Strong Acid: - as strong acid dissociates completely ($K_a > 1$) in water, the major species of all strong acids are H_3O^+ and their conjugate bases.



Example 5: List the major species $\text{H}_2\text{SO}_4(aq)$, and calculate its pH if it has a concentration of 1.00×10^{-5} M.



Since H_2SO_4 is a strong acid, the major species are: $\text{H}_3\text{O}^+(aq)$, $\text{HSO}_4^-(aq)$ and $\text{H}_2\text{O}(l)$.

(Even if H_2SO_4 is a diprotic acid, it dissociates one proton at a time. The conjugate base, HSO_4^- , is a weak acid. Weak acids dissociate differently than strong acids – next section.)

$$[\text{H}_3\text{O}^+] = [\text{H}_2\text{SO}_4]_0 = 1.00 \times 10^{-5} \text{ M}$$

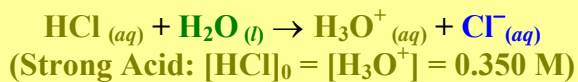
$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log (1.00 \times 10^{-5})$$

$$\text{pH} = 5.00$$

Example 6: Determine the $[\text{H}_3\text{O}^+]$ and/or $[\text{OH}^-]$ concentrations of the following solutions at 25°C.

a. $[\text{HCl}] = 0.350 \text{ mol/L}$

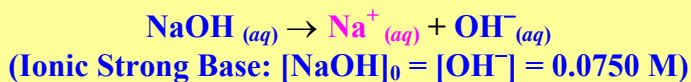


$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{0.350}$$

$$[\text{OH}^-] = 2.86 \times 10^{-14} \text{ mol/L}$$

b. $[\text{NaOH}] = 0.0750 \text{ M}$



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

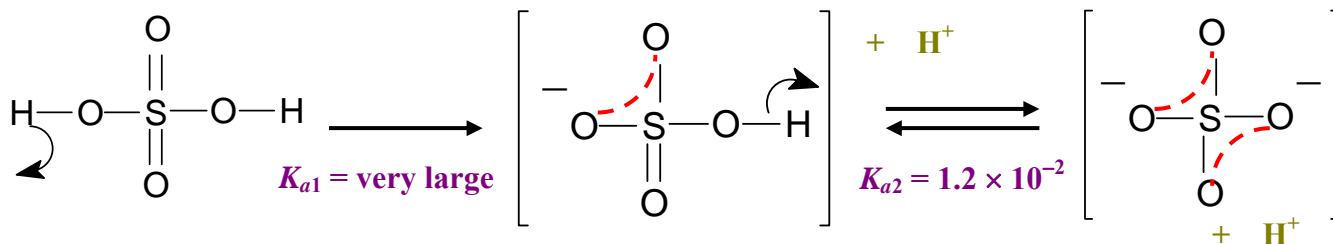
$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.0750}$$

$$[\text{H}_3\text{O}^+] = 1.33 \times 10^{-13} \text{ mol/L}$$

Oxoacids: - acids where the donating proton is attached to an oxygen atom.

- most acids are oxoacids because of the strong electronegativity of the oxygen atom, the hydrogen atom is more readily to leave as H^+ ion.

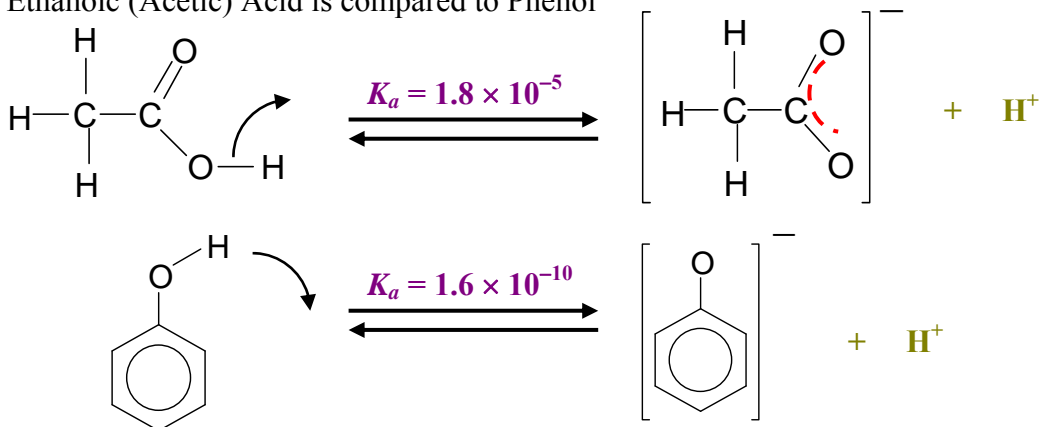
Example: Sulfuric Acid ($\text{H}_2\text{SO}_4_{(aq)}$), successively donate protons to finally reach $\text{SO}_4^{2-}_{(aq)}$



Organic Acids: - acids of organic compounds, commonly known as carboxylic acids.

- alcohols can also be considered as organic acids, but they are really weak. This is because carboxylic acids have double-bonded oxygen for resonance to occur, whereas alcohols only contain single-bonded oxygen.

Example: Ethanoic (Acetic) Acid is compared to Phenol



Assignment

15.3 pg. 687 #12 to 26

15.4 pg. 687–688 #27, 28, 30 to 38; pg. 690–692 #96, 112, 126, 140

15.5: Weak Acids and Acid Ionization Constants

Major Species of a Weak Acid: - since weak acid do not dissociate completely ($K_a < 1$) in water, the major species of all weak acids are their original form and water.

Procedure to calculate pH of Weak Acid Solutions:

1. List all the major species from all weak acids, including water.
2. Determine which species has the highest K_a . This is the Strongest Acid (SA) of the list.
3. Write the Brønsted-Lowry dissociation of this Strongest Acid.
4. Set up the **ICE Box** and the **equilibrium expression**.
5. If the original weak acid concentration, $[HA]_0$, is much larger than K_a , we can approximate by assuming $[HA]_{eq} = ([HA]_0 - x) \approx [HA]_0$. Thereby, simplifying the calculation. (**A general rule of thumb: if $[HA]_0 \geq 1000 \times K_a$, we can use the approximation.**)
6. Find the concentration of H_3O^+ .

7. Verify any approximation made by using the 5% rule. $\left(\frac{[H_3O^+]}{[HA]_0} \times 100\% \leq 5\% \right)$

Example 1: Determine the $[H_3O^+]$, and pH of 0.200 mol/L of HF (aq) ($K_a = 7.2 \times 10^{-4}$).

Major Species: HF (aq), H₂O (l)

Strongest Acid: HF ($K_a = 7.2 \times 10^{-4}$) (K_a for H₂O is 1.0×10^{-14})

HF is a weak acid and undergoes Brønsted-Lowry Dissociation.

	HF (aq)	H ₂ O (l)	H ₃ O ⁺ (aq)	F ⁻ (aq)
Initial	0.200 M		0	0
Change	-x		+x	+x
Equilibrium	(0.2 - x)		x	x

CANNOT use Approximation:

$$\frac{[HF]_0}{K_a} = \frac{0.200 \text{ M}}{7.2 \times 10^{-4}} = 277.8 < 1000$$

Have to use (0.2 - x) in the denominator

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} \quad 7.2 \times 10^{-4} = \frac{(x)(x)}{(0.2-x)} = \frac{x^2}{(0.2-x)}$$

$$7.2 \times 10^{-4} (0.2 - x) = x^2$$

$$(1.44 \times 10^{-4}) - (7.2 \times 10^{-4})x = x^2$$

$$x^2 + (7.2 \times 10^{-4})x - (1.44 \times 10^{-4}) = 0$$

(Quadratic Equation: Apply the Quadratic Formula!)

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \quad a = 1 \quad b = 7.2 \times 10^{-4} \quad c = -1.44 \times 10^{-4}$$

$$x = \frac{-(7.2 \times 10^{-4}) \pm \sqrt{(7.2 \times 10^{-4})^2 - 4(1)(-1.44 \times 10^{-4})}}{2(1)}$$

$$x = 0.0116453988$$

$$x = -0.0247307976 \text{ (omit negative } x)$$

$$[H_3O^+] = 0.0116 \text{ mol/L}$$

$$\text{pH} = -\log [H_3O^+]$$

$$\text{pH} = -\log(0.0116)$$

$$\text{pH} = 1.93$$

Verify that we could NOT use Approximation:

$$\frac{[H_3O^+]}{[HF]_0} \times 100\% = \frac{0.0116 \text{ M}}{0.200 \text{ M}} \times 100\% = 5.8\% > 5\%$$

Therefore, approximation would NOT be appropriate.

Percent Dissociation: - the amount of $[H_3O^+]$ dissociated from the original $[HA]_0$ expressed in percentage.
 - strong acids will have % dissociation = 100%.
 - weak acids will have % Dissociation < 100%

Percent Dissociation of Acids

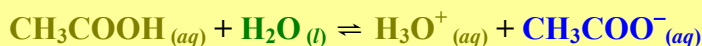
$$\% \text{ Dissociation} = \frac{[H_3O^+]}{[HA]_0} \times 100\%$$

Example 2: Determine the $[H_3O^+]$ and pH of an acid mixture of 0.500 mol/L HOBr (aq) ($K_a = 2 \times 10^{-9}$) of 0.200 mol/L of CH₃COOH (aq) ($K_a = 1.8 \times 10^{-5}$). Calculate the % dissociation of this acid mixture.

Major Species: HOBr (aq), CH₃COOH (aq), H₂O (l)

Strongest Acid: CH₃COOH ($K_a = 1.8 \times 10^{-5}$) (K_a for HOBr and H₂O are 2×10^{-9} and 1.0×10^{-14})

CH₃COOH is a weak acid and undergoes Brønsted-Lowry Dissociation.



	[CH ₃ COOH]	[H ₃ O ⁺]	[CH ₃ COO ⁻]
Initial	0.200 M	0	0
Change	-x	+x	+x
Equilibrium	(0.2 - x)	x	x

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} \quad 1.8 \times 10^{-5} = \frac{(x)(x)}{(0.2-x)} \approx \frac{x^2}{(0.2)}$$

$$1.8 \times 10^{-5} (0.2) \approx x^2$$

$$3.6 \times 10^{-6} \approx x^2$$

$$x \approx \sqrt{3.6 \times 10^{-6}}$$

$$x \approx 0.00190$$

CAN use Approximation:

$$\frac{[CH_3COOH]_0}{K_a} = \frac{0.200 \text{ M}}{1.8 \times 10^{-5}}$$

$$= 11111 \geq 1000$$

Use 0.2 in the denominator, because $(0.2 - x) \approx 0.2$ [x is so small compared to 0.2 M]

$$[H_3O^+] = 0.00190 \text{ mol/L}$$

$$\text{pH} = -\log [H_3O^+]$$

$$\text{pH} = -\log(0.00190)$$

$$\text{pH} = 2.72$$

Verify that we could use Approximation:

$$\frac{[H_3O^+]}{[CH_3COOH]_0} \times 100\% = \frac{0.00190 \text{ M}}{0.200 \text{ M}} \times 100\%$$

$$= 0.95\% \leq 5\%$$

Therefore, approximation would be appropriate.

$$\% \text{ Dissociation} = \frac{[H_3O^+]}{[CH_3COOH]_0} \times 100\%$$

$$\% \text{ Dissociation} = \frac{0.00190 \text{ M}}{0.200 \text{ M}} \times 100\%$$

$$\% \text{ Dissociation} = 0.95\%$$

Example 3: A 0.0500 mol/L of an unknown acid, HA, has a percent dissociation of 0.38%. What is the acid dissociation constant of this acid?

Major Species: HA (aq), H₂O (l)

Strongest Acid: HA ($K_a = ?$) (K_a for H₂O is 1.0×10^{-14} and has a 0.000 01 % dissociation)

HA is a weak acid and undergoes Brønsted-Lowry Dissociation.



	[HA]	[H ₃ O ⁺]	[A ⁻]
Initial	0.0500 M	0	0
Change	-1.9 × 10 ⁻⁴	+1.9 × 10 ⁻⁴	1.9 × 10 ⁻⁴
Equilibrium	0.04981 M	1.9 × 10 ⁻⁴ M	1.9 × 10 ⁻⁴ M

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(1.9 \times 10^{-4})(1.9 \times 10^{-4})}{(0.04981)}$$

$$\% \text{ Dissociation} = \frac{[H_3O^+]}{[HA]_0} \times 100\%$$

$$[H_3O^+] = (\% \text{Dissociation})[HA]_0 / 100\%$$

$$[H_3O^+] = (0.38\%)(0.0500 \text{ M}) / 100\%$$

$$[H_3O^+] = 1.9 \times 10^{-4} \text{ M} = [A^-]$$

$$K_a = 7.2 \times 10^{-7}$$

15.6 & 15.7: Weak Bases and Base Ionization Constants & The Relationship Between the Ionization and their Conjugate Bases

Strong Bases: - bases that **dissociate completely (100%) in water**.

- all **alkali bases (Group IA cations with OH⁻)** and some **alkaline bases (Group IIA cations with OH⁻)** are considered as strong bases because they are ionic compound that dissociates completely.

Examples: Strong Ionic Bases:

- a. Alkali Bases: LiOH_(aq), NaOH_(aq), KOH_(aq), RbOH_(aq), and CsOH_(aq)
(gives off 1 mole of OH⁻ when 1 mole of alkali base is dissolved)



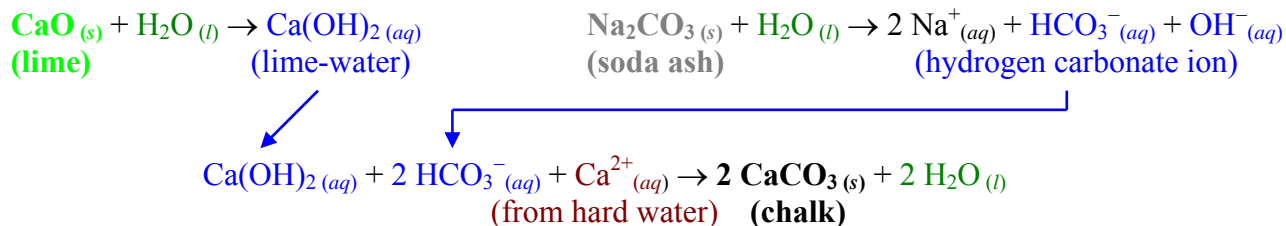
- b. Alkaline Bases: Ca(OH)_{2(aq)}, Ba(OH)_{2(aq)}, Sr(OH)_{2(aq)}
(gives off 2 moles of OH⁻ when 1 mole of alkaline base is dissolved)



Slaked Lime: - calcium hydroxide, Ca(OH)_{2(s)}; commonly refer to as lime-water when it is dissolved as Ca(OH)_{2(aq)}

- in reality, slaked lime do **not** dissolve that well in water. Its **dissociation constant (solubility product – K_{sp}) is 1.3 × 10⁻⁶** (more on K_{sp} in the next chapter). However, it is readily available and as such, it makes good bases in common laboratories.

Lime-Soda Process: - a process where **lime (CaO)** and **soda ash (Na₂CO₃)** are **added to water** in order to produce **chalk (CaCO₃)**.



Major Species of a Strong Base: - as alkali- and alkaline bases dissociate completely in water, **the major species of all strong bases are OH⁻ and water**.

- special care must be taken with **alkaline bases** as they **generate 2 moles of OH⁻ per 1 mole of solid dissolved**.



Example 1: List the major species $\text{NaOH}_{(aq)}$, and calculate its pH, pOH, $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ if it has a concentration of $1.00 \times 10^{-4} \text{ M}$.

$$\text{NaOH}_{(aq)} \rightarrow \text{Na}^+_{(aq)} + \text{OH}^-_{(aq)}$$

Since NaOH is a strong base, the major species are: $\text{OH}^-_{(aq)}$ and $\text{H}_2\text{O}_{(l)}$.

$$[\text{OH}^-] = [\text{NaOH}]_0 = 1.00 \times 10^{-4} \text{ M} \quad \text{pOH} = -\log [\text{OH}^-]$$

$$\text{pOH} = -\log (1.00 \times 10^{-4}) \quad \text{pOH} = 4.00$$

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 14 - (4.00)$$

$$\text{pH} = 10.00$$

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

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-10.00}$$

$$[\text{H}_3\text{O}^+] = 1.00 \times 10^{-10} \text{ M}$$

Example 2: List the major species $\text{Sr}(\text{OH})_2_{(aq)}$, and calculate its pH, pOH, $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ if it has a concentration of $1.00 \times 10^{-4} \text{ M}$.

$$\text{Sr}(\text{OH})_2_{(aq)} \rightarrow \text{Sr}^{2+}_{(aq)} + 2 \text{OH}^-_{(aq)}$$

Since $\text{Sr}(\text{OH})_2$ is a strong base, the major species are: $\text{OH}^-_{(aq)}$ and $\text{H}_2\text{O}_{(l)}$.

$$[\text{OH}^-] = 2 \times [\text{Sr}(\text{OH})_2]_0 = 2.00 \times 10^{-4} \text{ M} \quad \text{pOH} = -\log [\text{OH}^-]$$

$$\text{pOH} = -\log (2.00 \times 10^{-4}) \quad \text{pOH} = 3.70$$

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

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{2.00 \times 10^{-4}}$$

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

$$\text{pH} = -\log (5.00 \times 10^{-11})$$

$$[\text{H}_3\text{O}^+] = 5.00 \times 10^{-11} \text{ mol/L} \quad \text{pH} = 10.3$$

Weak Bases: - bases that dissociate LESS than 100% in water.

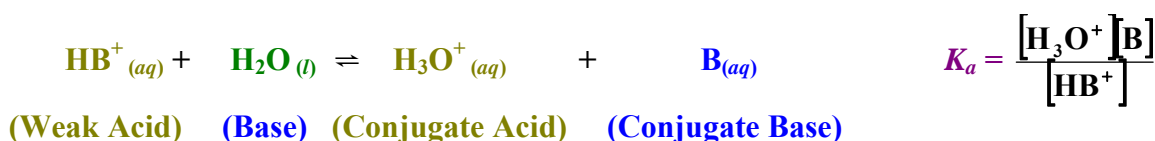
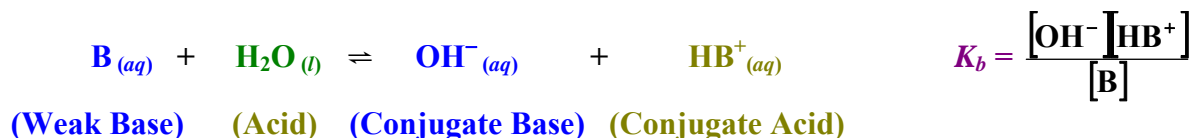
- when the base dissociation constant (K_b) is less than 1. ($K_b < 1$)
- the equilibrium position strongly favours the reactants.
- at equilibrium, the hydroxide concentration is much less than the original base concentration, $[\text{OH}^-]_{eq} < [\text{B}]_{eq}$ or $[\text{B}]_0 \approx [\text{B}]_{eq}$.
- the conjugate acid, HA^+ , of a weak base is itself a stronger weak acid (can easily donate protons to do the reverse reaction).



	[B]	[OH ⁻]	[HB ⁺]
Initial	x	0	0
Change	-y (where y << x)	+y	+y
Equilibrium	(x - y) ≈ x	y	y

Examples: Some Weak Bases: $\text{HOCCOO}^-_{(aq)}$, $\text{HSO}_3^-_{(aq)}$, $\text{HSO}_4^-_{(aq)}$, $\text{H}_2\text{PO}_4^-_{(aq)}$, $\text{NO}_2^-_{(aq)}$, $\text{HCO}_3^-_{(aq)}$, $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-_{(aq)}$, $\text{F}^-_{(aq)}$, $\text{HCOO}^-_{(aq)}$, $\text{C}_6\text{H}_7\text{O}_6^-_{(aq)}$, $\text{C}_6\text{H}_5\text{COO}^-_{(aq)}$, $\text{CO}_3^{2-}_{(aq)}$, $\text{CH}_3\text{COO}^-_{(aq)}$, $\text{HS}^-_{(aq)}$, $\text{OCl}^-_{(aq)}$, $\text{CN}^-_{(aq)}$, $\text{NH}_3_{(aq)}$, and $\text{NO}_3^-_{(aq)}$

Base Dissociation Constant (K_b): - the equilibrium constant of a dissociation of a weak base in water.
 - K_b of a weak base can be calculated from K_a of its conjugate acid and K_w .



$$K_a \times K_b = \frac{[\text{H}_3\text{O}^+][\text{B}]}{[\text{HB}^+]} \times \frac{[\text{OH}^-][\text{HB}^+]}{[\text{B}]} = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

Relationship between Conjugate Acid-Base Pair Dissociation Constants

$$K_w = K_a \times K_b$$

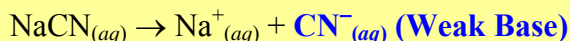
Major Species of a Weak Base: - since weak acid do not dissociate completely ($K_b < 1$) in water, the major species of all weak bases are their original form and water.

Procedure to calculate pH of Weak Base Solutions:

- List all the major species from all weak bases, including water.
- Calculate K_b from all weak bases using $K_w = K_a \times K_b$
- Determine which species has the highest K_b . This is the Strongest Base (SB) of the list.
- Write the Brønsted-Lowry dissociation of this Strongest Base
- Set up the **ICE Box** and the **equilibrium expression**.
- If the original weak base concentration, $[\text{B}]_0$, is much larger than K_b , we can approximate by assuming $[\text{B}]_{eq} = ([\text{B}]_0 - x) \approx [\text{B}]_0$. Thereby, simplifying the calculation. (**A general rule of thumb: if $[\text{B}]_0 \geq 1000 \times K_b$, we can use the approximation.**)
- Find the concentration of OH^- .
- Verify any approximation made by using the 5% rule. $\left(\frac{[\text{OH}^-]}{[\text{B}]_0} \times 100\% \leq 5\% \right)$

Example 3: Determine the $[\text{OH}^-]$, $[\text{H}_3\text{O}^+]$, pOH, pH and % dissociation of 0.200 mol/L of $\text{NaCN}_{(aq)}$. (K_a of $\text{HCN} = 6.2 \times 10^{-10}$).

NaCN dissociates completely in water:



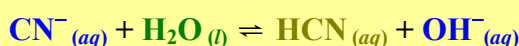
Major Species: $\text{CN}^-_{(aq)}$, $\text{H}_2\text{O}_{(l)}$

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{6.2 \times 10^{-10}} \quad K_b = 1.613 \times 10^{-5}$$

(taking a few more decimal places to avoid round off errors)

Strongest Base: CN^- ($K_b = 1.613 \times 10^{-5}$) (K_b for H_2O is $K_w = 1.0 \times 10^{-14}$)

CN^- is a weak base and undergoes Brønsted-Lowry Dissociation.



	$[\text{CN}^-]$	$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$
Initial	0.200 M	0	0
Change	-x	+x	+x
Equilibrium	(0.2 - x)	x	x

$$K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]} \quad 1.613 \times 10^{-5} = \frac{(x)(x)}{(0.2-x)} \approx \frac{x^2}{(0.2)}$$

$$1.613 \times 10^{-5} (0.2) \approx x^2$$

$$3.226 \times 10^{-6} \approx x^2$$

$$x \approx \sqrt{3.226 \times 10^{-6}}$$

$$x \approx 0.0018$$

CAN use Approximation:

$$\frac{[\text{CN}^-]_0}{K_b} = \frac{0.200 \text{ M}}{1.613 \times 10^{-5}}$$

$$= 12399 \geq 1000$$

Use 0.2 in the denominator, because $(0.2 - x) \approx 0.2$ [x is so small compared to 0.2 M]

$$[\text{OH}^-] = 0.0018 \text{ mol/L} = 1.8 \text{ mmol/L}$$

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

$$\text{pOH} = -\log(0.0018)$$

$$\text{pOH} = 2.75$$

Verify that we could use Approximation:

$$\frac{[\text{OH}^-]}{[\text{CN}^-]_0} \times 100\% = \frac{0.0018 \text{ M}}{0.200 \text{ M}} \times 100\%$$

$$= 0.90\% \leq 5\%$$

$$\% \text{ Dissociation} = \frac{[\text{OH}^-]}{[\text{CN}^-]_0} \times 100\%$$

$$\% \text{ Dissociation} = \frac{0.0018 \text{ M}}{0.200 \text{ M}} \times 100\%$$

$$\% \text{ Dissociation} = 0.90\%$$

Therefore, approximation would be appropriate.

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 14 - (2.75)$$

$$\text{pH} = 11.25$$

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

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-11.25}$$

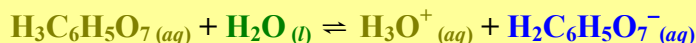
$$[\text{H}_3\text{O}^+] = 5.6 \times 10^{-12} \text{ M}$$

Assignment

15.5 pg. 688 #39 to 50; pg. 690–691 #98 and 124
15.6 & 15.7 pg. 688–689 #51 to 58 pg. 691 #122

15.8: Diprotic and Polyprotic Acids**Polyprotic Acids:** - acids that can donate more than one protons.

- this includes all diprotic and **triprotic acids** (acids that can **donate three protons**).
- **polyprotic acids dissociate one proton at a time**. Each successive proton donation has its own K_a , which gets smaller until the last proton is donated. ($K_{a1} > K_{a2} > K_{a3} > \dots$).
- the **intermediates (conjugate bases of each dissociation except the last one) are themselves acids**. Thus, *diprotic acid can generate one amphoteric species*, and *triprotic acid can generate two amphoteric species*.
- **except for sulfuric acid (because K_{a1} is very large and K_{a2} is small), all polyprotic acids' $[\text{H}_3\text{O}^+]$ are calculated from their K_{a1}** because successive proton donation from smaller subsequent K_a do not amount to any significant increase in the $[\text{H}_3\text{O}^+]$.

Examples: Some Polyprotic Acids:a. Diprotic Acids: $\text{H}_2\text{SO}_4(aq)$, $\text{HOOC}\text{COOH}(aq)$, $\text{H}_2\text{SO}_3(aq)$, $\text{H}_2\text{C}_6\text{H}_6\text{O}_6(aq)$, $\text{H}_2\text{CO}_3(aq)$, $\text{H}_2\text{S}(aq)$ Diprotic Amphoteric Intermediates: $\text{HSO}_4^-(aq)$, $\text{HOOC}\text{COO}^-(aq)$, $\text{HSO}_3^-(aq)$, $\text{HC}_6\text{H}_6\text{O}_6^-(aq)$, $\text{HCO}_3^-(aq)$, $\text{HS}^-(aq)$ b. Triprotic Acids: $\text{H}_3\text{PO}_4(aq)$, $\text{H}_3\text{AsO}_4(aq)$, $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$, $\text{H}_3\text{BO}_3(aq)$ Triprotic Amphoteric Intermediates: $\text{H}_2\text{PO}_4^-(aq)$, $\text{HPO}_4^{2-}(aq)$, $\text{H}_2\text{AsO}_4^-(aq)$, $\text{HASO}_4^{2-}(aq)$, $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(aq)$, $\text{HC}_6\text{H}_5\text{O}_7^{2-}(aq)$, $\text{H}_2\text{BO}_3^-(aq)$, $\text{HBO}_3^{2-}(aq)$ **Example 1:** Determine the pH of 2.00 M of citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$) and the concentrations of $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$, $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(aq)$, $\text{HC}_6\text{H}_5\text{O}_7^{2-}(aq)$, and $\text{C}_6\text{H}_5\text{O}_7^{3-}(aq)$. The acid dissociation constants are $K_{a1} = 8.4 \times 10^{-4}$, $K_{a2} = 1.8 \times 10^{-5}$, and $K_{a3} = 4.0 \times 10^{-6}$.**Major Species:** $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$, $\text{H}_2\text{O}(l)$ **Strongest Acid:** $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$ ($K_{a1} = 8.4 \times 10^{-4}$) $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$ is a weak acid (Brønsted-Lowry Dissociation).

	$[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]$	$[\text{H}_3\text{O}^+]$	$[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]$
Initial	2.00 M	0	0
Change	-x	+x	+x
Equilibrium	(2-x)	x	x

CAN use Approximation:

$$\frac{[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]_0}{K_{a1}} = \frac{2.00 \text{ M}}{8.4 \times 10^{-4}} = 2381 \geq 1000$$

Use 2 in the denominator, because $(2-x) \approx 2$ [x is so small compared to 2.00 M]

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]}{[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]}$$

$$8.4 \times 10^{-4} = \frac{(x)(x)}{(2-x)} \approx \frac{x^2}{2}$$

$$8.4 \times 10^{-4} (2) \approx x^2$$

$$0.00168 \approx x^2$$

$$x \approx \sqrt{0.00168}$$

$$x \approx 0.0410$$

$$[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]_{eq} = 2.00 \text{ M} - 0.0410 \text{ M}$$

$$[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]_{eq} = 1.96 \text{ M}$$

$$[\text{H}_3\text{O}^+] = 0.0410 \text{ mol/L}$$

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

$$\text{pH} = -\log(0.0410)$$

$$\text{pH} = 1.39$$

(after 1st proton donation)**Verify that we could use Approximation:**

$$\frac{[\text{H}_3\text{O}^+]}{[\text{H}_3\text{C}_6\text{H}_5\text{O}_7]_0} \times 100\% = \frac{0.0410 \text{ M}}{2.00 \text{ M}} \times 100\%$$

$$= 2.05\% \leq 5\%$$

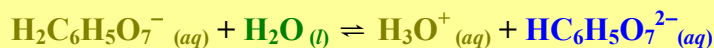
Therefore, approximation would be appropriate.

Next, we have to calculate $[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]$ after the first proton donation.

Major Species: $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-$ (aq), H_2O (l)

Strongest Acid: $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-$ (aq) ($K_{a2} = 1.8 \times 10^{-5}$)

$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-$ is a weak acid (Brønsted-Lowry Dissociation).



	$[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]$	$[\text{H}_3\text{O}^+]$	$[\text{HC}_6\text{H}_5\text{O}_7^{2-}]$
Initial	$x = 0.0410 \text{ M}$	0.0410 M	0
Change	$-y$	$+y$	$+y$
Equilibrium	$(0.041 - y)$	$(0.041 + y)$	y

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HC}_6\text{H}_5\text{O}_7^{2-}]}{[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]} \quad 1.8 \times 10^{-5} = \frac{(0.041 + y)(y)}{(0.041 - y)} \approx \frac{(0.041)y}{(0.041)}$$

Verify that we could use Approximation:

$$\frac{[\text{H}_3\text{O}^+]}{[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]_0} \times 100\% = \frac{1.8 \times 10^{-5} \text{ M}}{0.0410 \text{ M}} \times 100\% = 0.44\% \leq 5\%$$

$$y \approx 1.8 \times 10^{-5}$$

Therefore, approximation would be appropriate.

$$\text{New } [\text{H}_3\text{O}^+] = 0.0410 \text{ M} + 1.8 \times 10^{-5} \text{ M}$$

$$\text{New pH} = -\log [\text{H}_3\text{O}^+] = -\log (0.041018)$$

$$\text{New } [\text{H}_3\text{O}^+] = 0.041018 \text{ M}$$

$$\text{New pH} = 1.39 \text{ (after second proton donation)}$$

$$[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]_{\text{eq}} = 0.041 \text{ M} - 1.8 \times 10^{-5} \text{ M}$$

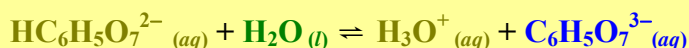
$$[\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-]_{\text{eq}} = 0.040982 \text{ M}$$

Finally, we have to calculate $[\text{HC}_6\text{H}_5\text{O}_7^{2-}]$ and $[\text{C}_6\text{H}_5\text{O}_7^{3-}]$ after the last proton donation.

Major Species: $\text{HC}_6\text{H}_5\text{O}_7^{2-}$ (aq), H_2O (l)

Strongest Acid: $\text{HC}_6\text{H}_5\text{O}_7^{2-}$ (aq) ($K_{a3} = 4.0 \times 10^{-6}$)

$\text{HC}_6\text{H}_5\text{O}_7^{2-}$ is a weak acid (Brønsted-Lowry Dissociation).



	$[\text{HC}_6\text{H}_5\text{O}_7^{2-}]$	$[\text{H}_3\text{O}^+]$	$[\text{C}_6\text{H}_5\text{O}_7^{3-}]$
Initial	$y = 1.8 \times 10^{-5} \text{ M}$	0.041018 M	0
Change	$-z$	$+z$	$+z$
Equilibrium	$(1.8 \times 10^{-5} - z)$	$(0.041018 + z)$	z

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_6\text{H}_5\text{O}_7^{3-}]}{[\text{HC}_6\text{H}_5\text{O}_7^{2-}]} \quad 4.0 \times 10^{-6} = \frac{(0.041018 + z)(z)}{(1.8 \times 10^{-5} - z)}$$

$$0 = \frac{(0.041018 + z)(z)}{(1.8 \times 10^{-5} - z)} - 4.0 \times 10^{-6}$$

CANNOT use Approximation:

$$\frac{[\text{HC}_6\text{H}_5\text{O}_7^{2-}]_0}{K_{a3}} = \frac{1.8 \times 10^{-5} \text{ M}}{4.0 \times 10^{-6}} = 4.5 < 1000$$

```

solve((0.041018+x)
)/((1.8E-5-x)-4E
-6,x,0,(0.18E-5
))
1.755155695E-9
  
```

$$z = 1.76 \times 10^{-9}$$

$$\text{Final } [\text{H}_3\text{O}^+] = 0.041018 \text{ M} + 1.76 \times 10^{-9} \text{ M}$$

$$\text{Final pH} = -\log [\text{H}_3\text{O}^+] = -\log(0.0410180018)$$

$$\text{Final } [\text{H}_3\text{O}^+] = 0.0410180018 \text{ M}$$

$$\text{Final pH} = 1.39$$

(no change from the pH at K_{a1})

$$[\text{HC}_6\text{H}_5\text{O}_7^{2-}]_{\text{eq}} = 1.8 \times 10^{-5} \text{ M} - 1.76 \times 10^{-9} \text{ M}$$

$$[\text{HC}_6\text{H}_5\text{O}_7^{2-}]_{\text{eq}} = 1.80 \times 10^{-5} \text{ M}$$

$$[\text{C}_6\text{H}_5\text{O}_7^{3-}]_{\text{eq}} = 1.76 \times 10^{-9} \text{ M}$$

Example 2: Determine the pH of 0.0500 M of sulfuric acid ($\text{H}_2\text{SO}_4(aq)$) and the concentrations of $\text{HSO}_4^-(aq)$, and $\text{SO}_4^{2-}(aq)$. The acid dissociation constant is $K_{a2} = 1.2 \times 10^{-2}$.



Since H_2SO_4 is a strong acid, the major species are: $\text{H}_3\text{O}^+(aq)$, $\text{HSO}_4^-(aq)$ and $\text{H}_2\text{O}(l)$.

$$[\text{H}_3\text{O}^+] = [\text{H}_2\text{SO}_4]_0 = 0.0500 \text{ M}$$

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

$$\text{pH} = -\log (0.0500)$$

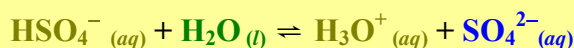
$$\text{pH} = 1.30 \text{ (after } K_{a1}\text{)}$$

Next, we have to calculate $[\text{HSO}_4^-]$, $[\text{SO}_4^{2-}]$ and the final pH after the last proton donation.

Major Species: $\text{HSO}_4^-(aq)$, $\text{H}_2\text{O}(l)$

Strongest Acid: $\text{HSO}_4^-(aq)$ ($K_{a2} = 1.2 \times 10^{-2}$)

HSO_4^- is a weak acid (Brønsted-Lowry Dissociation).



	$[\text{HSO}_4^-]$	$[\text{H}_3\text{O}^+]$	$[\text{SO}_4^{2-}]$
Initial	0.0500 M	0.0500 M	0
Change	-x	+x	+x
Equilibrium	(0.05 - x)	(0.05 + x)	x

CANNOT use Approximation:

$$\frac{[\text{HSO}_4^-]_0}{K_{a2}} = \frac{0.0500 \text{ M}}{1.2 \times 10^{-2}} = 4.167 < 1000$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{SO}_4^{2-}]}{[\text{HSO}_4^-]}$$

$$1.2 \times 10^{-2} = \frac{(0.05 + x)(x)}{(0.05 - x)}$$

$$0 = \frac{(0.05 + x)(x)}{(0.05 - x)} - 1.2 \times 10^{-2}$$

$$x = 0.00851$$

```
solve((0.05+x)*x/
(0.05-x)-1.2E-2,
x,0,(0.05))
.0085094925
```

Final $[\text{H}_3\text{O}^+] = 0.0500 \text{ M}$ (from K_{a1}) + 0.00851 M (from K_{a2})

Final $[\text{H}_3\text{O}^+] = 0.0585 \text{ M}$
(a significant change from before)

$$[\text{HSO}_4^-]_{eq} = 0.0500 \text{ M} - 0.00851 \text{ M}$$

$$\text{Final pH} = -\log [\text{H}_3\text{O}^+] = -\log (0.0585)$$

$$[\text{HSO}_4^-]_{eq} = 0.0415 \text{ M}$$

$$[\text{SO}_4^{2-}]_{eq} = 0.00585 \text{ M}$$

Final pH = 1.23
(different from 1.30 of the pH at K_{a1})

From the last two examples, we can see that sulfuric acid ($\text{H}_2\text{SO}_4(aq)$) requires the calculation of $[\text{H}_3\text{O}^+]$ in both steps of the dissociation. Other polyprotic acids, such as $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$ only require the first dissociation step to calculate the $[\text{H}_3\text{O}^+]$.

Assignment

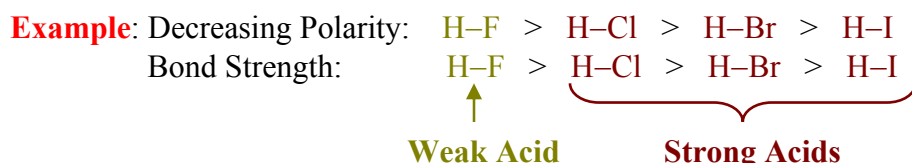
15.8 pg. 689 #59 to 64; pg. 691 #118 and 120

15.9: Molecular Structure and the Strength of Acids**Structural Factors that affect Acid Properties**

1. **Polarity:** - in general, the **more polar the intramolecular bond between hydrogen and the adjacent atom**, the more likely hydrogen be donated and becoming a **stronger Brønsted-Lowry acid**. (C–H bonds have very little polarity.)

Hydrohalic Acids (H–X): - any acids that contain a halogen atom as adjacent atom to the donating hydrogen atom.

2. **Bond Strength:** - **the stronger the bond strength** (more exothermic ΔH_f) **means weaker Brønsted-Lowry acid**. This is because the stronger the bond strength, the more energy it will be needed to break the bond between hydrogen and the adjacent atom. Therefore, the proton is less likely to be donated.



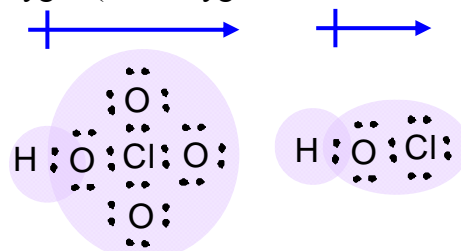
Despite the high polarity, HF is a weak acid because of its high bond strength. This is also due to the fact that F is in the second orbital and its protons have a more effective “pull” on the bonding electrons. Thereby, keeping the hydrogen atom from donating as a H^+ ion.

Oxoacids (H–O–X): - any acids that contain an oxygen atom as adjacent atom to the donating hydrogen atom.
 - the other side of this oxygen atom is bonded by a non-metal atom.

Examples: Some oxoacids: $\text{HClO}_3(aq)$, $\text{H}_2\text{SO}_4(aq)$, $\text{HNO}_2(aq)$

3. **Electron Density of Oxoacids:** - the **more oxygen atoms that an oxoacid has within a series**, the **stronger the acid**. This is due to the higher electron density for the oxoacid with higher number of oxygen (each oxygen atom can provide two lone-pairs).

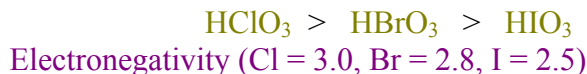
Example: Decreasing Strength of oxoacid series:



Higher Electron Density with More Oxygen Atoms in Oxoacids makes Stronger Acids

4. **Electronegativity:** - the **higher the electronegativity** of the oxoacid's central atom (H–O–X) with the same number of oxygen atoms, the **stronger the acid**. This is because higher electronegativity atom draws bonding electrons closer itself, leaving a weaker bond between the oxygen and the hydrogen. The result is the increasing tendency for the hydrogen to donate (making it a stronger acid).

Example: Decreasing Strength of various oxoacid series with the same number of oxygen atoms:



5. **Resonance**: - if the conjugate base has a resonance structure after the donation of a proton, the stronger is the original acid.

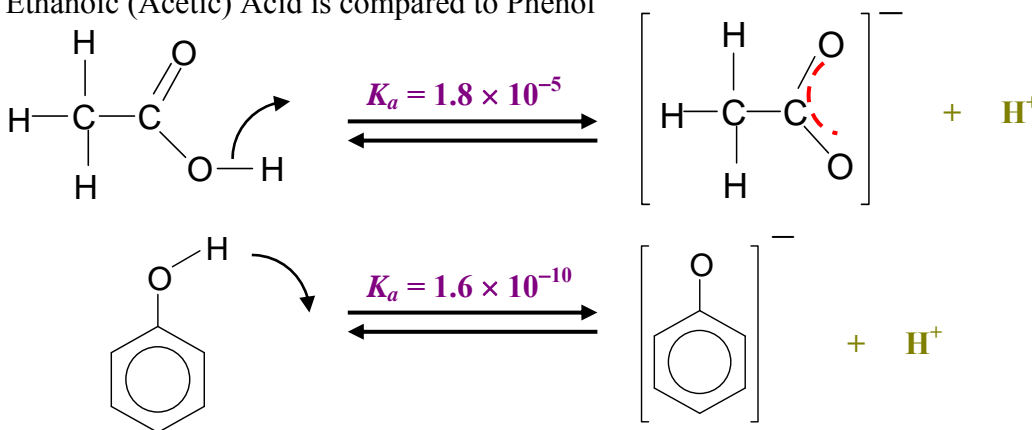
Carboxylic Acids (R-COO-H): - any organic acids that contain the carboxyl group(s)

- the conjugate base has a stable resonance structure, thereby makes proton donation of the acid possible.

Examples: Some carboxylic acids: $\text{HCOOH}_{(aq)}$, $\text{CH}_3\text{COOH}_{(aq)}$, $\text{HOOC-COOH}_{(aq)}$

- alcohols can also be considered as organic acids, but they are really weak (so weak that we don't really think of them as acids). This is because carboxylic acids have double-bonded oxygen for resonance to occur, whereas alcohols only contain single-bonded oxygen.

Example: Ethanoic (Acetic) Acid is compared to Phenol



15.10: Acid-Base Properties of Salts

Salts: - ionic compounds that might dissociate in water.

1. **Neutral Salts**: - when the **Cation comes from a Strong Base** and the **Anion is the Conjugate-Base of a Strong Acid**.
 - **no effect on pH; if dissolve in pure water, pH will remain at 7.**

Examples: Some Neutral Salts:

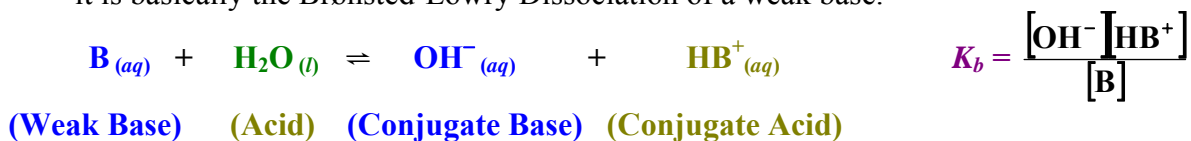
- $\text{KNO}_3_{(aq)}$ (K^+ can be from a strong base – $\text{KOH}_{(aq)}$; NO_3^- is the conjugate-base of a strong acid – $\text{HNO}_3_{(aq)}$)
- $\text{NaCl}_{(aq)}$ (Na^+ can be from a strong base – $\text{NaOH}_{(aq)}$; Cl^- is the conjugate-base a strong acid – $\text{HCl}_{(aq)}$)

2. **Basic Salts**: - when the **Cation comes from a Strong Base** and the **Anion is the Conjugate-Base of a Weak Acid**.
 - **pH will increase; if dissolve in pure water, pH > 7 (Basic).**

Examples: Some Basic Salts:

- $\text{NaCH}_3\text{COO}_{(aq)}$ (Na^+ can be from a strong base – $\text{NaOH}_{(aq)}$; CH_3COO^- is the conjugate-base of a weak acid – $\text{CH}_3\text{COOH}_{(aq)}$)
- $\text{KF}_{(aq)}$ (K^+ can be from a strong base – $\text{KOH}_{(aq)}$; F^- is the conjugate base of a weak acid – $\text{HF}_{(aq)}$)

Hydrolysis: - the reaction of a base and a water to form a conjugate acid and OH^- .
 - it is basically the Brønsted-Lowry Dissociation of a weak base.



Percent Hydrolysis: - the amount of $[\text{OH}^-]$ dissociated from the original $[\text{B}]_0$ expressed in percentage.
 - strong bases will have % Hydrolysis = 100%.
 - weak bases will have % Hydrolysis < 100%

Percent Hydrolysis of Bases

$$\% \text{ Hydrolysis} = \frac{[\text{OH}^-]}{[\text{B}]_0} \times 100\%$$

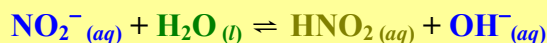
Example 1: Determine the pH and the percent hydrolysis of 0.235 M of sodium nitrite. The acid dissociation constant for nitrous acid is $K_a = 4.0 \times 10^{-4}$.

Sodium nitrite (NaNO_2) dissociates completely in H_2O : $\text{NaNO}_{2(s)} \rightarrow \text{Na}^+_{(aq)} + \text{NO}_2^-_{(aq)}$ (Weak Base)

Major Species: $\text{NO}_2^-_{(aq)}$, $\text{H}_2\text{O}_{(l)}$ $K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{4.0 \times 10^{-4}} \quad K_b = 2.5 \times 10^{-11}$

Strongest Base: NO_2^- ($K_b = 2.5 \times 10^{-11}$) (K_b for H_2O is $K_w = 1.0 \times 10^{-14}$)

NO_2^- is a weak base and undergoes Brønsted-Lowry Dissociation.



	$[\text{NO}_2^-]$	$[\text{HNO}_2]$	$[\text{OH}^-]$
Initial	0.235 M	0	0
Change	-x	+x	+x
Equilibrium	(0.235 - x)	x	x

$$K_b = \frac{[\text{HNO}_2][\text{OH}^-]}{[\text{NO}_2^-]} \quad 2.5 \times 10^{-11} = \frac{(x)(x)}{(0.235 - x)} \approx \frac{x^2}{(0.235)}$$

$$2.5 \times 10^{-11} (0.235) \approx x^2$$

$$5.875 \times 10^{-12} \approx x^2$$

$$x \approx \sqrt{5.875 \times 10^{-12}}$$

$$x = [\text{OH}^-] \approx 2.4 \times 10^{-6} \text{ M}$$

CAN use Approximation:

$$\frac{[\text{NO}_2^-]_0}{K_b} = \frac{0.235 \text{ M}}{2.5 \times 10^{-11}} = 9.4 \times 10^9 \geq 1000$$

Use 0.235 in the denominator, because $(0.235 - x) \approx 0.235$ [x is so small compared to 0.235 M]

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

$$\text{pOH} = -\log(2.4 \times 10^{-6})$$

$$\text{pOH} = 5.62$$

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 14 - (5.62)$$

$$\text{pH} = 8.38$$

$$\% \text{ Hydrolysis} = \frac{[\text{OH}^-]}{[\text{NO}_2^-]_0} \times 100\% = \frac{2.4 \times 10^{-6} \text{ M}}{0.235 \text{ M}} \times 100\%$$

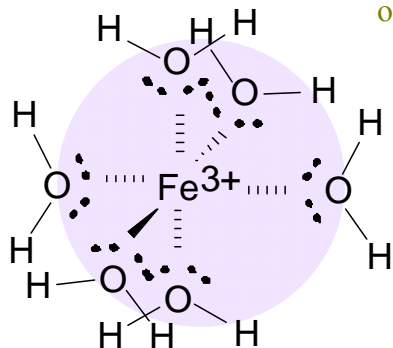
$$\% \text{ Hydrolysis} = 0.00102\%$$

As the small K_b suggests, the $[\text{OH}^-]$ is very small compared to $[\text{NO}_2^-]_0$. Therefore, % hydrolysis is also very small as a result.

3. **Acidic Salts:** - when the **Cation is the Conjugate-Acid of a Weak Base** and the **Anion is the Conjugate-Base of a Strong Acid**.
 - when the **Cation is a Metal Ion that is Highly Charged** (Hydrated Complex Cations – $M(H_2O)_n^{m+}$) and the **Anion is the Conjugate-Base of a Strong Acid**.
 - **pH will decrease; if dissolve in pure water, pH < 7 (Acidic).**

Examples: Some Acidic Salts:

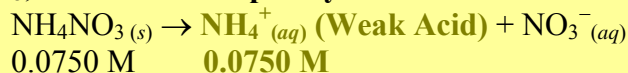
- $NH_4ClO_4(aq)$ (NH_4^+ is the conjugate-acid of a weak base – $NH_3(aq)$; ClO_4^- is the conjugate-base of a strong acid – $HClO_4(aq)$)
- $AlCl_3(aq)$ (Al^{3+} can be form a hydrated complex ion $[Al(H_2O)_6]^{3+}(aq)$; Cl^- is the conjugate-base of a strong acid – $HCl(aq)$)
- $Fe(NO_3)_3(aq)$ (Fe^{3+} can be form a hydrated complex ion $[Fe(H_2O)_6]^{3+}(aq)$; NO_3^- is the conjugate-base of a strong acid – $HNO_3(aq)$)
- $CoI_3(aq)$ (Co^{3+} can be form a hydrated complex ion $[Co(H_2O)_6]^{3+}(aq)$; I^- is the conjugate-base of a strong acid – $HI(aq)$)



- **hydrated metal ions that are highly charged attracts the lone pair of water molecules**, allowing the **electron density to shift towards the metal ion center** and hence, **hydrogen from the exterior water molecules can be donated more readily**. In effect, it makes the hydrated metal ions acidic. Some hydrated metal ions are Al^{3+} , Cr^{3+} , Co^{3+} , Fe^{3+} , Bi^{3+} and Be^{2+} .

Example 2: Determine the pH of 0.0750 M of ammonium nitrate. The base dissociation constant for $NH_3(aq)$ is $K_b = 1.8 \times 10^{-5}$.

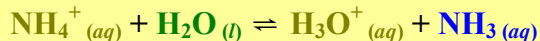
Ammonium nitrate, NH_4NO_3 , dissociates completely in water:



Major Species: $NH_4^+(aq)$, $H_2O(l)$ $K_a = \frac{K_w}{K_b} = \frac{1.00 \times 10^{-14}}{1.8 \times 10^{-5}} \quad K_a = 5.556 \times 10^{-10}$

Strongest Acid: NH_4^+ ($K_a = 5.556 \times 10^{-10}$) (K_a for H_2O is $K_w = 1.0 \times 10^{-14}$)

NH_4^+ is a weak acid and undergoes Brønsted-Lowry Dissociation.



	$[NH_4^+]$	$[H_3O^+]$	$[NH_3]$
Initial	0.0750 M	0	0
Change	-x	+x	+x
Equilibrium	(0.075 - x)	x	x

$$K_a = \frac{[H_3O^+][NH_3]}{[NH_4^+]} \quad 5.556 \times 10^{-10} = \frac{(x)(x)}{(0.075 - x)} \approx \frac{x^2}{(0.075)}$$

CAN use Approximation:

$$\frac{[NH_4^+]_0}{K_a} = \frac{0.0750 \text{ M}}{5.556 \times 10^{-10}} = 1.35 \times 10^8 \geq 1000$$

Use 0.075 in the denominator, because $(0.075 - x) \approx 0.075$ [x is so small compared to 0.075 M]

Verify for Using Approximation: $5.556 \times 10^{-10} (0.075) \approx x^2$
 $4.167 \times 10^{-11} \approx x^2$

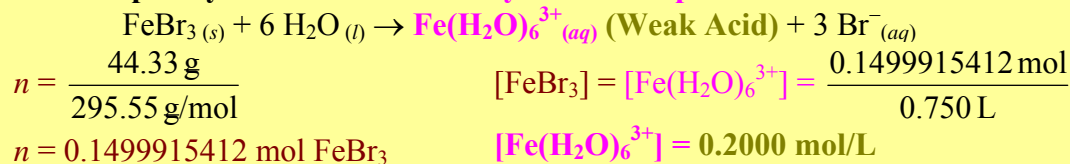
$$\frac{[H_3O^+]}{[NH_4^+]_0} \times 100\% = \frac{6.455 \times 10^{-6} \text{ M}}{0.0750 \text{ M}} \times 100\% \quad x \approx \sqrt{4.167 \times 10^{-11}} \quad \text{pH} = -\log [H_3O^+]$$

$$= 0.0086\% \leq 5\% \quad \text{Therefore, approximation would be appropriate.} \quad x = [H_3O^+] \approx 6.455 \times 10^{-6} \text{ M} \quad \text{pH} = -\log (6.455 \times 10^{-6})$$

pH = 5.19

Example 3: Determine the pH of 44.33 g of iron (III) bromide (FeBr_3) dissolved in 750 mL of water. The acid dissociation constant for $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$ is $K_a = 8.3 \times 10^{-3}$.

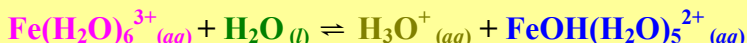
FeBr_3 dissociates completely in water and form hydrated complex ion:



Major Species: $\text{Fe}(\text{H}_2\text{O})_6^{3+}(aq)$, $\text{H}_2\text{O}(l)$

Strongest Acid: $\text{Fe}(\text{H}_2\text{O})_6^{3+}$ ($K_a = 8.3 \times 10^{-3}$)

$\text{Fe}(\text{H}_2\text{O})_6^{3+}$ is a weak acid and undergoes Brønsted-Lowry Dissociation.



	$[\text{Fe}(\text{H}_2\text{O})_6^{3+}]$	$[\text{H}_3\text{O}^+]$	$[\text{FeOH}(\text{H}_2\text{O})_5^{2+}]$
Initial	0.200 M	0	0
Change	-x	+x	+x
Equilibrium	(0.2 - x)	x	x

CANNOT use Approximation:

$$\frac{[\text{Fe}(\text{H}_2\text{O})_6^{3+}]_0}{K_a} = \frac{0.200 \text{ M}}{8.3 \times 10^{-3}} = 24.1 < 1000$$

```

solve(X^2/(.2-X)-
8.3E-3, X, 0, .2
)
.036803907

```

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{FeOH}(\text{H}_2\text{O})_5^{2+}]}{[\text{Fe}(\text{H}_2\text{O})_6^{3+}]}$$

$$8.3 \times 10^{-3} = \frac{(x)(x)}{(0.2-x)} = \frac{x^2}{(0.2-x)}$$

$$0 = \frac{x^2}{(0.2-x)} - 8.3 \times 10^{-3}$$

$$x = [\text{H}_3\text{O}^+] = 0.037 \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (0.037)$$

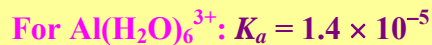
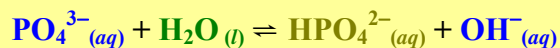
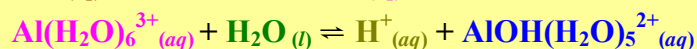
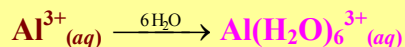
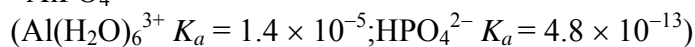
pH = 1.43

Salts that Contain both Conjugate-Acid and Conjugate-Base:

- If the **conjugate-acid ion has a greater K_a than the K_b of the conjugate-base ion**, then the solution will be **acidic** ($\text{pH} < 7$ when $K_a \text{ of Conj-Acid} > K_b \text{ of Conj-Base}$).
- If the **conjugate-base ion has a greater K_b than the K_a of the conjugate-acid ion**, then the solution will be **basic** ($\text{pH} > 7$ when $K_b \text{ of Conj-Base} > K_a \text{ of Conj-Acid}$).
- When **both K_a from the conjugate-acid ion is equal to the K_b of the conjugate-base ion**, then the solution will be **neutral** ($\text{pH} = 7$ when $K_a \text{ of Conj-Acid} = K_b \text{ of Conj-Base}$).

Example 4: Classify the following salts as acid, base or neutral.

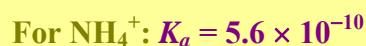
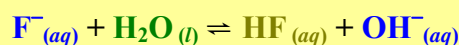
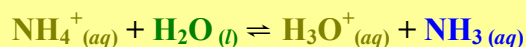
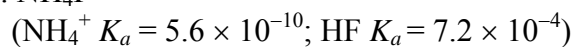
a. AlPO_4



$$\text{For } \text{PO}_4^{3-}: K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.8 \times 10^{-13}} = 0.021$$

Since K_b of $\text{PO}_4^{3-} > K_a$ of $\text{Al}(\text{H}_2\text{O})_6^{3+}$,
aluminum phosphate is a Basic Salt.

b. NH_4F



$$\text{For } \text{F}^-: K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{7.2 \times 10^{-4}} = 1.4 \times 10^{-11}$$

Since K_a of $\text{NH}_4^+ > K_b$ of F^- ,
ammonium fluoride is an Acidic Salt.

Assignment

15.9 pg. 689 #67 to 70; pg. 691 #130

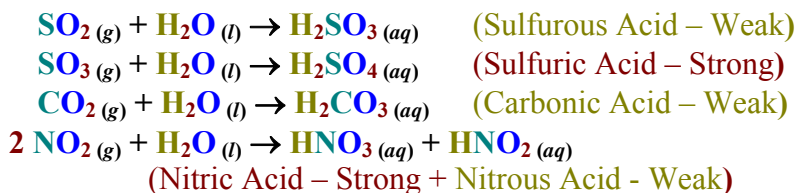
15.10 pg. 689 #75 to 82; pg. 691 #106 and 128

15.11: Acid-Base Properties of Oxides and Hydroxides

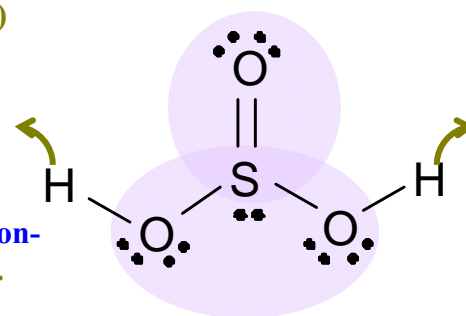
Acidic Oxides: - **molecular oxides (XO_n)** where **X is a non-metal atom** have a **tendency to form acids** when dissolved in water.

- this is due to the fact that the **non-metal atoms have a high electronegativity**. Thereby, they draw the electrons of the adjacent oxygen closer themselves. Hence, the **O–H bond** within X–O–H **becomes so weak** that **hydrogen ion (H^+) can be easily donated**.

Example: The problem of acid rain is created when various molecular oxides from the exhaust of vehicles and fossil-fuel power generator (coal or natural gas) react with water in the atmosphere.



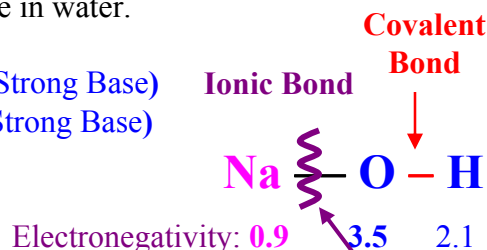
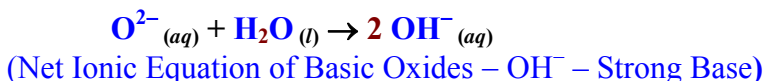
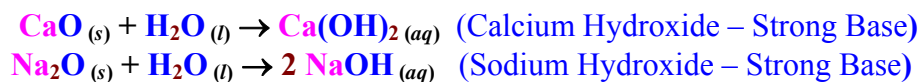
Higher Electron Density Around the Central Non-Metal Atom makes H^+ more ready to donate.



Basic Oxides: - **ionic oxides (M_mO_n)** where **M is a metal atom** have a **tendency to form bases** when dissolved in water.

- this is due to the fact that the metal atoms have small electronegativity. Thereby, the electrons of the adjacent oxygen. Hence, the **O–H bond** within M–O–H **becomes so strong** that **hydroxide ion (OH^-) can be easily produced** (Arrhenius Base).

Example: Some Metal Oxides that can form a base when dissolve in water.



Bigger Difference in Electronegativities with Ionic Bonds (bond will likely break between the Metal atom and the Oxygen atom) – producing OH^- ion as a result.

Amphoteric Hydroxides: - metal hydroxides (non-alkali metals and non-alkaline metals – except $Be(OH)_2$) that can behave as acids or bases under different environments.
- some amphoteric hydroxides are $Be(OH)_2$, $Al(OH)_3$, $Sn(OH)_2$, $Pb(OH)_2$, $Cr(OH)_2$, $Cu(OH)_2$, $Zn(OH)_2$, and $Cd(OH)_2$.



15.12: The Lewis Acids and Bases

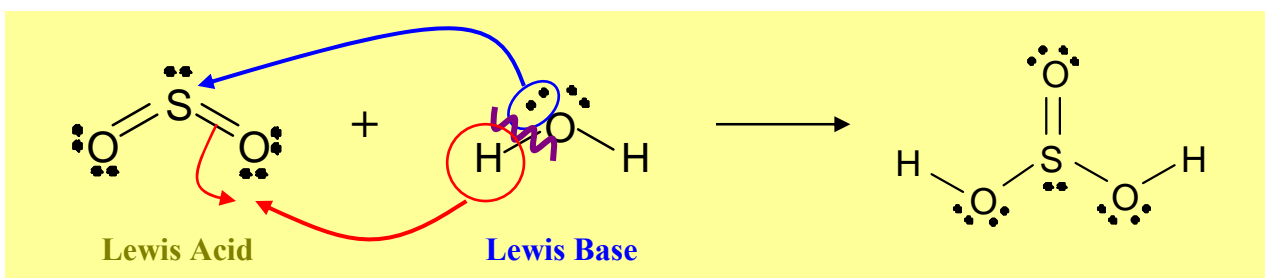
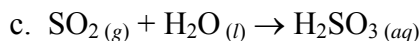
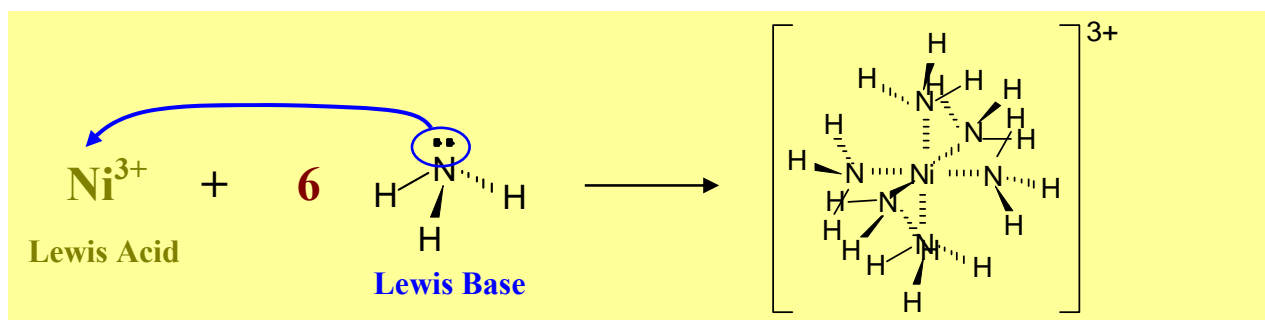
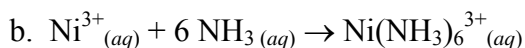
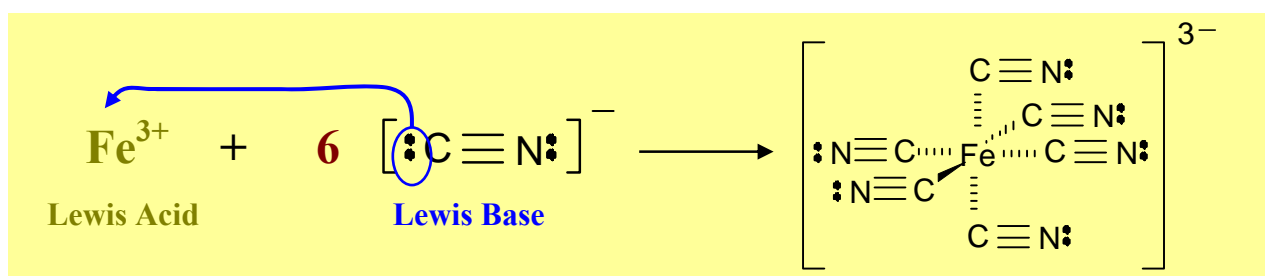
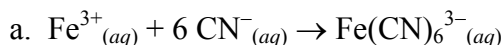
Lewis Acid-Base Model: - using **electron pair** instead of **proton**, we can define acids that do not have a proton and bases that do not have a hydroxide ion.

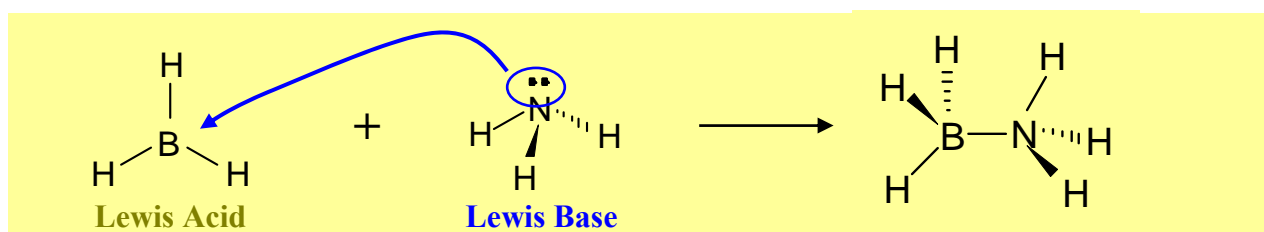
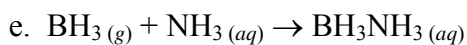
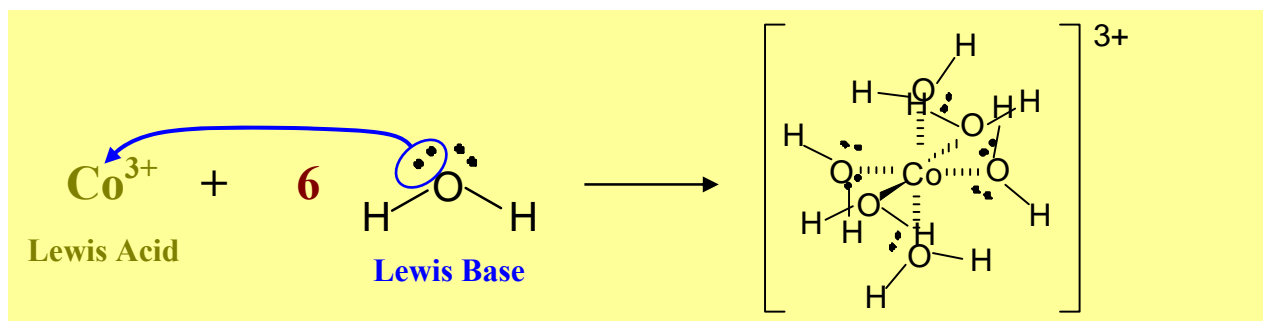
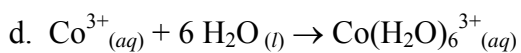
- especially suitable to explain metal hydrated cation as acid, or how metal oxides can turn into a base.

Lewis Acid: - a substance that **ACCEPTS an Electron Pair (Lone-Pair)**.

Lewis Base: - a substance that **DONATES an Electron Pair (Lone-Pair)**.

Example 1: Draw the Lewis diagrams for the reactants and products. Identify the Lewis acids and bases.





Assignment

15.11 pg. 690 #83 to 88

15.12 pg. 690 #89 to 94; pg. 692 #140

Chapter 16: Acid-Base Equilibria and Solubility Equilibria**16.1: Homogeneous Versus Heterogeneous Solution Equilibria**

In acids and bases chemistry as well as solubility of ionic products, the equilibria involve are usually heterogeneous (where the phases of chemical species in the equilibrium are not in the same phases). Hence, great care is required when handling problems of these equilibria.

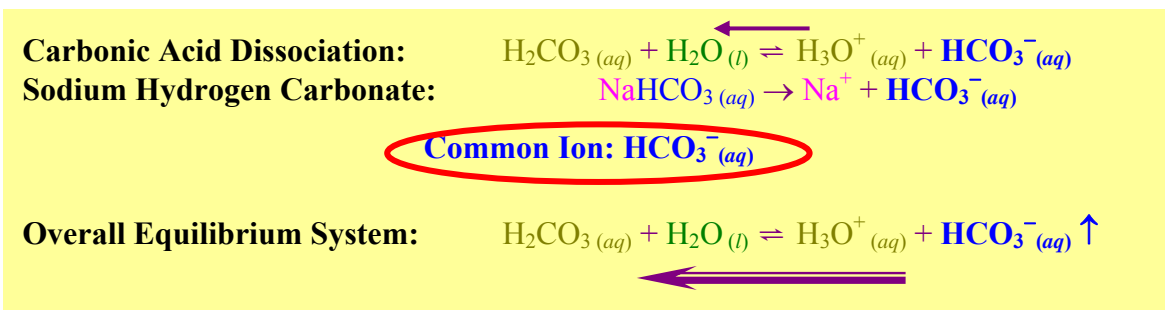
16.2 & 16.3: The Common Ion Effect & Buffer Solutions

Common Ion: - the ion that is present in two separate solutions as they are added together.

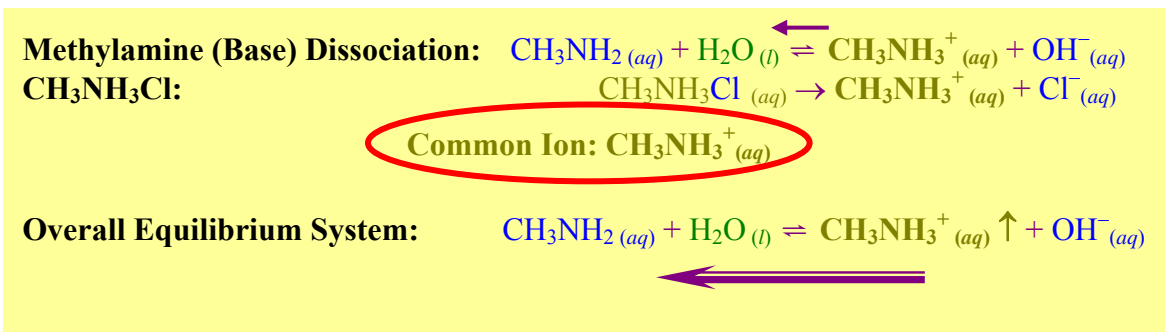
Common Ion Effect: - the shifting of the equilibrium, as outlined by Le Châtelier's principle, due to the addition or presence of a common ion in system.

Example 1: Identify the common ion in the following solutions.

- a. 0.0250 M of carbonic acid with 0.0300 M of sodium hydrogen carbonate.



- b. 0.500 M of methylamine with 0.450 M of $\text{CH}_3\text{NH}_3\text{Cl}$

**pH Calculation Involving Common Ion:**

1. Determine the **Initial Concentration of ALL Major Species** in the equilibrium.
2. Set up the **ICE Box and the Equilibrium Expression Equating it to the Equilibrium Constant**.
3. Calculate the $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$ and the **pH**.

pH Calculations Involving Buffered Solutions and any Subsequent Addition of H⁺ or OH⁻ Amounts:

1. Determine the **initial concentration of all major species** in the equilibrium.
2. Set up the **ICE Box and the Equilibrium Expression Equating it to the Equilibrium Constant**.
3. Calculate the **[H₃O⁺]** or **[OH⁻]** and the **pH** of the **original buffered solution**.
4. **Write out the Complete Reaction for the ADDITION of H⁺ or OH⁻ into the system**. Because we are adding a strong acid or a strong base, we will **Treat the Reaction as One-Way**. Hence, we can **Apply Regular Stoichiometric Principle** to calculate the moles of each major species.
5. Again, set up the **ICE Box and the Equilibrium Expression Equating it to the Equilibrium Constant**.
6. Calculate the **[H₃O⁺]** or **[OH⁻]** and the **pH** of the **revised buffered solution**.

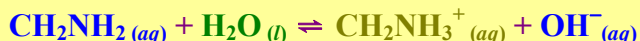
Example 3: Calculate the pH of a 1.00 L buffered solution consisting of 0.500 M of methylamine ($K_b = 4.38 \times 10^{-4}$) with 0.450 M of CH₃NH₃Cl when:

- 0.0200 mol of KOH is added to it.
- 3.00 mL of 0.750 M of HNO₃(aq) is added to it.
- Contrast the pH's of the above two additions against the same additions to 1.00 L of water.

First, we have to figure out the [OH⁻] and pH for the buffered solution system.

Methylamine (Base) Dissociation: $\text{CH}_2\text{NH}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_2\text{NH}_3^+(aq) + \text{OH}^-(aq)$

CH₃NH₃Cl: $\text{CH}_3\text{NH}_3\text{Cl}(aq) \rightarrow \text{CH}_2\text{NH}_3^+(aq) + \text{Cl}^-(aq)$
 0.450 M 0.450 M



	[CH ₂ NH ₂]	[CH ₂ NH ₃ ⁺]	[OH ⁻]
Initial	0.500 M	0.450 M	0
Change	-x	+x	+x
Equilibrium	(0.5 - x)	(0.45 + x)	x

CAN use Approximation:

$$\frac{[\text{CH}_2\text{NH}_2]_0}{K_b} = \frac{0.500 \text{ M}}{4.38 \times 10^{-4}} = 1142 \geq 1000$$

$$K_b = \frac{[\text{OH}^-][\text{CH}_2\text{NH}_3^+]}{[\text{CH}_2\text{NH}_2]} \quad 4.38 \times 10^{-4} = \frac{(x)(0.45 + x)}{(0.5 - x)} \approx \frac{x(0.45)}{(0.5)}$$

$$4.38 \times 10^{-4} \frac{(0.5)}{(0.45)} \approx x$$

$$x \approx 4.87 \times 10^{-4}$$

Use 0.45 in the numerator, because (0.45 + x) ≈ 0.45 [x is so small compared to 0.45 M].
 Use 0.5 in the denominator, because (0.5 - x) ≈ 0.5 [x is so small compared to 0.5 M].

Verify that we could use Approximation:

$$\frac{[\text{OH}^-]}{[\text{CH}_2\text{NH}_2]_0} \times 100\% = \frac{4.87 \times 10^{-4} \text{ M}}{0.500 \text{ M}} \times 100\%$$

$$= 0.0973\% \leq 5\%$$

Therefore, approximation would be appropriate.

$$[\text{OH}^-] = 4.87 \times 10^{-4} \text{ mol/L}$$

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

$$\text{pOH} = -\log(4.87 \times 10^{-4})$$

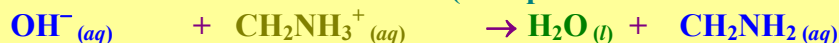
$$\text{pOH} = 3.31$$

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 14 - 3.31$$

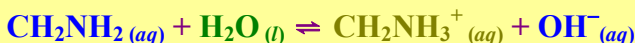
$$\text{pH} = 10.69$$

a. 0.0200 mol of KOH is added to the buffered solution (Complete Rxn between OH^- & CH_2NH_3^+).



	n of OH^-	n of CH_2NH_3^+	n of CH_2NH_2
Before	0.0200 mol (Limiting)	0.450 mol	0.500 mol
Change	-0.0200 mol	-0.0200 mol	+0.0200 mol
After	0	0.430 mol	0.520 mol

Recalculate equilibrium concentrations using NEW $[\text{CH}_2\text{NH}_3^+]_0$ and $[\text{CH}_2\text{NH}_2]_0$



	$[\text{CH}_2\text{NH}_2]$	$[\text{CH}_2\text{NH}_3^+]$	$[\text{OH}^-]$
Initial	0.520 M	0.430 M	0
Change	-y	+y	+y
Equilibrium	(0.52 - y)	(0.43 + y)	y

CAN use Approximation:

$$\frac{[\text{CH}_2\text{NH}_2]_0}{K_b} = \frac{0.520 \text{ M}}{4.38 \times 10^{-4}} = 1187 \geq 1000$$

Use 0.43 in the numerator, because $(0.43 + y) \approx 0.43$ [y is so small compared to 0.43 M].

Use 0.52 in the denominator, because $(0.52 - y) \approx 0.52$ [y is so small compared to 0.52 M].

$$K_b = \frac{[\text{OH}^-][\text{CH}_2\text{NH}_3^+]}{[\text{CH}_2\text{NH}_2]} \quad 4.38 \times 10^{-4} \approx \frac{y(0.43)}{(0.52)}$$

$$4.38 \times 10^{-4} = \frac{(y)(0.43 + y)}{(0.52 - y)} \quad 4.38 \times 10^{-4} \frac{(0.52)}{(0.43)} \approx y$$

$$y \approx 5.30 \times 10^{-4}$$

Verify that we could use Approximation:

$$\frac{[\text{OH}^-]}{[\text{CH}_2\text{NH}_2]_0} \times 100\% = \frac{5.30 \times 10^{-4} \text{ M}}{0.520 \text{ M}} \times 100\%$$

$$= 0.102\% \leq 5\% \text{ (Appropriate Approximation)}$$

$$[\text{OH}^-] = 5.30 \times 10^{-4} \text{ mol/L}$$

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

$$\text{pOH} = -\log(5.30 \times 10^{-4})$$

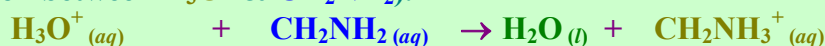
$$\text{pOH} = 3.28$$

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 14 - 3.28$$

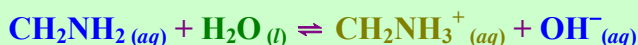
$$\text{pH} = 10.72$$

b. 3.00 mL of 0.750 M (0.00225 mol) of $\text{HNO}_3 (aq)$ is added to the buffered solution (Complete Reaction between H_3O^+ & CH_2NH_2).



	n of H_3O^+	n of CH_2NH_2	n of CH_2NH_3^+
Before	0.00225 mol (Limiting)	0.500 mol	0.450 mol
Change	-0.00225 mol	-0.00225 mol	+0.00225 mol
After	0	0.49775 mol	0.45225 mol

Recalculate equilibrium concentrations using NEW $[\text{CH}_2\text{NH}_2]_0$ and $[\text{CH}_2\text{NH}_3^+]_0$ (neglect Δ volume)



CAN use Approximation:

$$\frac{[\text{CH}_2\text{NH}_2]_0}{K_b} = \frac{0.520 \text{ M}}{4.38 \times 10^{-4}} = 1187 \geq 1000$$

Use 0.45225 in the numerator, because $(0.45225 + z) \approx 0.45225$ [z is so small compared to 0.45225 M].

Use 0.49775 in the denominator, because $(0.49775 - z) \approx 0.49775$ [z is so small compared to 0.49775 M].

$$K_b = \frac{[\text{OH}^-][\text{CH}_2\text{NH}_3^+]}{[\text{CH}_2\text{NH}_2]} \quad 4.38 \times 10^{-4} \approx \frac{z(0.45225)}{(0.49775)}$$

$$4.38 \times 10^{-4} = \frac{(z)(0.45225 + z)}{(0.49775 - z)} \quad 4.38 \times 10^{-4} \frac{(0.49775)}{(0.45225)} \approx z$$

$$z \approx 4.82 \times 10^{-4}$$

Verify that we could use Approximation:

$$\frac{[\text{OH}^-]}{[\text{CH}_2\text{NH}_2]_0} \times 100\% = \frac{4.82 \times 10^{-4} \text{ M}}{0.49775 \text{ M}} \times 100\% = 0.0968\% \leq 5\% \text{ (Appropriate Approximation)}$$

$$[\text{OH}^-] = 4.82 \times 10^{-4} \text{ mol/L}$$

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ \text{pOH} &= -\log(4.82 \times 10^{-4}) \end{aligned}$$

$$\text{pOH} = 3.32$$

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 14 - 3.28$$

$$\text{pH} = 10.68$$

c. Contrast the pH's of the above two additions against the same additions to 1.00 L of water.

i. 0.0200 mol of KOH is added to 1.00 L of water

$$\text{pOH} = -\log [\text{OH}^-] \quad \text{pH} = 14 - \text{pOH}$$

$$\text{pOH} = -\log(0.0200) \quad \text{pH} = 14 - 1.70$$

$$\text{pOH} = 1.70$$

$$\text{pH} = 12.30$$

Adding the 0.0200 mol of KOH to buffered solution.

pH changes from 10.69 to 10.72

ii. 3.00 mL of 0.750 M (0.00225 mol) of $\text{HNO}_3(aq)$ is added to 1.00 L of water (neglect Δ volume).

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

$$\text{pH} = -\log(0.00225)$$

$$\text{pH} = 2.65$$

Adding the 0.00225 mol of HNO_3 to buffered solution.

pH changes from 10.69 to 10.68

Acid/Conjugate Base Ratio $\left(\frac{[\text{HA}]}{[\text{A}^-]}\right)$: together with K_a , it governs the $[\text{H}_3\text{O}^+]$ of a buffered solution.

- it can be used to adjust the pH range of the buffered solution.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \implies [\text{H}_3\text{O}^+] = K_a \left(\frac{[\text{HA}]}{[\text{A}^-]}\right)$$

(Assumes that $[\text{HA}]_0 \approx ([\text{HA}]_{eq} - x) \approx [\text{HA}]_{eq}$ and $[\text{A}^-]_0 \approx ([\text{HA}]_{eq} + x) \approx [\text{A}^-]_{eq}$ because K_a is usually small)

$$[\text{H}_3\text{O}^+] = K_a \left(\frac{[\text{HA}]}{[\text{A}^-]}\right) \quad \text{(Relationship of Conjugate Base/Acid Ratio of Buffered Solution)}$$

$$-\log[\text{H}_3\text{O}^+] = -\log \left[K_a \left(\frac{[\text{HA}]}{[\text{A}^-]}\right) \right] \quad \text{(Log Both sides)}$$

$$-\log[\text{H}_3\text{O}^+] = -\log K_a - \log \left(\frac{[\text{HA}]}{[\text{A}^-]}\right) \quad \text{(Apply Logarithmic Law: } \log(MN) = \log M + \log N)$$

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]}\right) \quad \text{(Apply Logarithmic Law: } \log(M^{-1}) = -\log M)$$

Henderson-Hasselbalch Equation

For a Buffered Solution with small K_a , assuming $[\text{HA}]_0 \approx [\text{HA}]_{eq}$ and $[\text{A}^-]_0 \approx [\text{A}^-]_{eq}$, it has a pH of:

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]}\right) \quad \text{(For Basic Buffered Solution } K_a = \frac{K_w}{K_b})$$

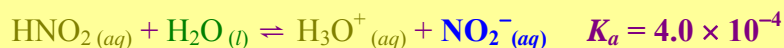
$$\text{p}K_a = -\log K_a$$

$[\text{A}^-]$ = initial [Conjugate Base]

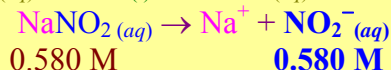
$[\text{HA}]$ = initial [Acid]

Example 4: Calculate the pH of a 1.00 L buffered solution consisting of 0.650 M of nitrous acid ($K_a = 4.0 \times 10^{-4}$) with 0.580 M of sodium nitrite solution. What is the new pH after 0.125 mol of $\text{Ba}(\text{OH})_2$ is added to it?

Nitrous Acid Dissociation:



Sodium Nitrite:



For the buffered solution we can use the Henderson-Hasselbalch Equation.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{NO}_2^-]}{[\text{HNO}_2]} \right)$$

$$\text{pH} = -\log K_a + \log \left(\frac{[\text{NO}_2^-]}{[\text{HNO}_2]} \right) = -\log(4.0 \times 10^{-4}) + \log \left(\frac{0.580 \text{ M}}{0.650 \text{ M}} \right)$$

Buffered Solution pH = 3.35

CAN use Approximation:

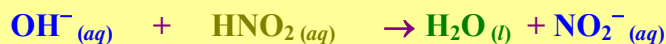
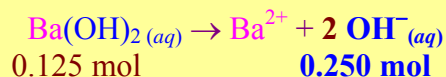
$$\frac{[\text{HNO}_2]_0}{K_a} = \frac{0.650 \text{ M}}{4.0 \times 10^{-4}}$$

$$= 1625 \geq 1000$$

Calculate new pH after 0.125 mol of $\text{Ba}(\text{OH})_2$ is added to the Buffered Solution.

(Complete Reaction between OH^- & HNO_2).

Barium Hydroxide:



	<i>n</i> of OH^-	<i>n</i> of HNO_2	<i>n</i> of NO_2^-
Before	0.250 mol (Limiting)	0.650 mol	0.580 mol
Change	- 0.250 mol	- 0.250 mol	+ 0.250 mol
After	0	0.400 mol	0.830 mol

CAN use Approximation:

$$\frac{[\text{HNO}_2]_0}{K_a} = \frac{0.400 \text{ M}}{4.0 \times 10^{-4}}$$

$$= 1000 \geq 1000$$

$$[\text{HNO}_2]_{eq} \approx [\text{HNO}_2]_0 = \frac{0.400 \text{ mol}}{1.00 \text{ L}} = 0.400 \text{ M}$$

$$[\text{NO}_2^-]_{eq} \approx [\text{NO}_2^-]_0 = \frac{0.830 \text{ mol}}{1.00 \text{ L}} = 0.830 \text{ M}$$

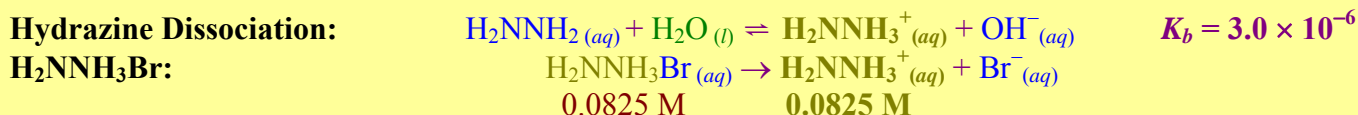
For the buffered solution with $\text{Ba}(\text{OH})_2$ added, we can again use the Henderson-Hasselbalch Equation.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{NO}_2^-]}{[\text{HNO}_2]} \right) = -\log K_a + \log \left(\frac{[\text{NO}_2^-]}{[\text{HNO}_2]} \right)$$

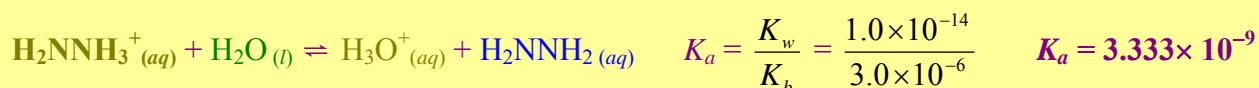
$$\text{pH} = -\log(4.0 \times 10^{-4}) + \log \left(\frac{0.830 \text{ M}}{0.400 \text{ M}} \right)$$

pH (Buffered Solution with $\text{Ba}(\text{OH})_2$ added) = 3.71

Example 5: Calculate the pH of a 1.00 L buffered solution consisting of 0.0750 M of hydrazine, H_2NNH_2 (aq), ($K_b = 3.0 \times 10^{-6}$) with 0.0825 M of $\text{H}_2\text{NNH}_3\text{Br}$ solution. What is the new pH after 15.0 mL of 1.00 mol/L HBr is added to it (neglect any volume changes)?



We need to write the acid dissociation reaction and obtain K_a .



For the buffered solution, we can use the Henderson-Hasselbalch Equation.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{H}_2\text{NNH}_2]}{[\text{H}_2\text{NNH}_3^+]} \right) = -\log K_a + \log \left(\frac{[\text{H}_2\text{NNH}_2]}{[\text{H}_2\text{NNH}_3^+]} \right)$$

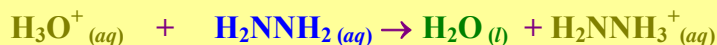
$$\text{pH} = -\log(3.333 \times 10^{-9}) + \log \left(\frac{0.0750 \text{ M}}{0.0825 \text{ M}} \right)$$

CAN use Approximation:

$$\frac{[\text{H}_2\text{NNH}_3^+]_0}{K_a} = \frac{0.0825 \text{ M}}{3.333 \times 10^{-9}} = 2.48 \times 10^7 \geq 1000$$

Buffered Solution pH = 8.44

Calculate new pH after 15.0 mL of 1.00 mol/L (0.0150 mol) HBr is added to the Buffered Solution. (Complete Reaction between H_3O^+ & H_2NNH_2).



	n of H_3O^+	n of H_2NNH_2	n of H_2NNH_3^+
Before	0.0150 mol (Limiting)	0.0750 mol	0.0825 mol
Change	-0.0150 mol	-0.0150 mol	+0.0150 mol
After	0	0.0600 mol	0.0975 mol

CAN use Approximation:

$$\frac{[\text{HNO}_2]_0}{K_a} = \frac{0.0975 \text{ M}}{3.333 \times 10^{-9}} = 2.93 \times 10^7 \geq 1000$$

$$[\text{H}_2\text{NNH}_3^+]_{eq} \approx [\text{H}_2\text{NNH}_3^+]_0 = \frac{0.0975 \text{ mol}}{1.00 \text{ L}} = 0.0975 \text{ M}$$

$$[\text{H}_2\text{NNH}_2]_{eq} \approx [\text{H}_2\text{NNH}_2]_0 = \frac{0.0600 \text{ mol}}{1.00 \text{ L}} = 0.0600 \text{ M}$$

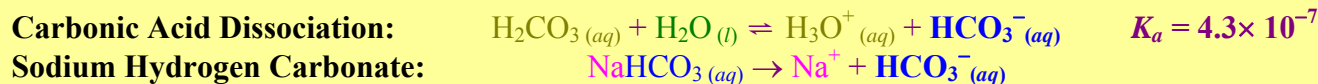
For the buffered solution with HBr added, we can again use the Henderson-Hasselbalch Equation.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{H}_2\text{NNH}_2]}{[\text{H}_2\text{NNH}_3^+]} \right) = -\log K_a + \log \left(\frac{[\text{H}_2\text{NNH}_2]}{[\text{H}_2\text{NNH}_3^+]} \right)$$

$$\text{pH} = -\log(3.333 \times 10^{-9}) + \log \left(\frac{0.0600 \text{ M}}{0.09750 \text{ M}} \right)$$

pH (Buffered Solution with HBr added) = 8.27

Example 6: Devise a procedure to make a buffered solution using carbonic acid ($K_{a1} = 4.3 \times 10^{-7}$ and $K_{a2} = 5.6 \times 10^{-11}$) and sodium hydrogen carbonate that is in the range of $\text{pH} = 7.50$.



For the buffered solution, we can use the Henderson-Hasselbalch Equation to find $\left(\frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}\right)$.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} \right)$$

$$\text{pH} - \text{p}K_a = \log \left(\frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} \right)$$

$$7.50 - [-\log(4.3 \times 10^{-7})] = \log \left(\frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} \right)$$

$$\log \left(\frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} \right) = 1.133468456$$

$$10^{1.133468456} = \left(\frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} \right)$$

$$\left(\frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} \right) = 13.59779394$$

One way to make this $\text{HCO}_3^-/\text{H}_2\text{CO}_3$ buffered solution ($\text{pH} = 7.50$) is to use **0.136 M** of $\text{NaHCO}_3(aq)$ with **0.0100 M** of $\text{H}_2\text{CO}_3(aq)$. The conjugate base/acid ratio would be $\frac{0.136 \text{ M}}{0.0100 \text{ M}} \approx 13.59779394$.

Buffering Capacity: - the amounts of H^+ or OH^- a buffered solution can handle before there is a significant change in pH.

- a large buffering capacity means it can absorb large amount of H^+ or OH^- before there is a notable change in pH.

- depends very much on the conjugate base/acid ratio $\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$.

- the optimal buffering occurs when $[\text{A}^-] = [\text{HA}]$ and they are relatively large compared to the amounts of H^+ or OH^- added.

Optimal Buffering Capacity

For an Optimal Buffering Capacity, a Buffered Solution should have large $[\text{HA}] = \text{large } [\text{A}^-]$.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right) \quad \text{pH} = \text{p}K_a \quad (\text{for Best Buffering Capacity})$$

Therefore, the desired pH for the buffered solution should be the SAME as the $\text{p}K_a$ of the weak acid, HA.

$$[\text{C}_2\text{H}_5\text{COOH}]_{eq} \approx [\text{C}_2\text{H}_5\text{COOH}]_0 = \frac{0.0150 \text{ mol}}{1.00 \text{ L}} = 0.0150 \text{ M}$$

$$[\text{C}_2\text{H}_5\text{COO}^-]_{eq} \approx [\text{C}_2\text{H}_5\text{COO}^-]_0 = \frac{0.0450 \text{ mol}}{1.00 \text{ L}} = 0.0450 \text{ M}$$

CAN use Approximation:

$$\frac{[\text{C}_2\text{H}_5\text{COOH}]_0}{K_a} = \frac{0.0150 \text{ M}}{1.3 \times 10^{-5}} = 1154 \geq 1000$$

For the buffered solution with NaOH added, we can again use the Henderson-Hasselbalch Equation.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{C}_2\text{H}_5\text{COO}^-]}{[\text{C}_2\text{H}_5\text{COOH}]} \right) = -\log K_a + \log \left(\frac{[\text{C}_2\text{H}_5\text{COO}^-]}{[\text{C}_2\text{H}_5\text{COOH}]} \right)$$

$$\text{pH} = -\log(1.3 \times 10^{-5}) + \log \left(\frac{0.0450 \text{ M}}{0.0150 \text{ M}} \right)$$

pH changed from 4.89 to 5.36

$$\text{pH (0.0300 M Buffered Solution with NaOH added)} = 5.36$$

Calculate new pH after 0.0150 mol of NaOH is added to the 3.00 M of $\text{NaC}_2\text{H}_5\text{COO}_{(aq)}/\text{C}_2\text{H}_5\text{COOH}_{(aq)}$ Buffered Solution. (Complete Reaction between OH^- & $\text{C}_2\text{H}_5\text{COOH}$).

	n of OH^-	n of $\text{C}_2\text{H}_5\text{COOH}$	n of $\text{C}_2\text{H}_5\text{COO}^-$
Before	0.0150 mol (Limiting)	3.00 mol	3.00 mol
Change	-0.0150 mol	-0.0150 mol	+0.0150 mol
After	0	2.985 mol	3.015 mol

$$[\text{C}_2\text{H}_5\text{COOH}]_{eq} \approx [\text{C}_2\text{H}_5\text{COOH}]_0 = \frac{2.985 \text{ mol}}{1.00 \text{ L}} = 2.985 \text{ M}$$

$$[\text{C}_2\text{H}_5\text{COO}^-]_{eq} \approx [\text{C}_2\text{H}_5\text{COO}^-]_0 = \frac{3.015 \text{ mol}}{1.00 \text{ L}} = 3.015 \text{ M}$$

CAN use Approximation:

$$\frac{[\text{C}_2\text{H}_5\text{COOH}]_0}{K_a} = \frac{2.985 \text{ M}}{1.3 \times 10^{-5}} = 229615 \geq 1000$$

For the buffered solution with NaOH added, we can again use the Henderson-Hasselbalch Equation.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{C}_2\text{H}_5\text{COO}^-]}{[\text{C}_2\text{H}_5\text{COOH}]} \right) = -\log K_a + \log \left(\frac{[\text{C}_2\text{H}_5\text{COO}^-]}{[\text{C}_2\text{H}_5\text{COOH}]} \right)$$

$$\text{pH} = -\log(1.3 \times 10^{-5}) + \log \left(\frac{3.015 \text{ M}}{2.985 \text{ M}} \right)$$

pH stayed at 4.89 from before

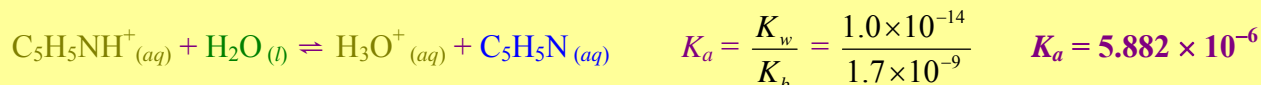
$$\text{pH (3.00 M Buffered Solution with NaOH added)} = 4.89$$

Hence, when $[\text{A}^-] = [\text{HA}]$ and if both concentrations are large, then we achieve optimal buffering capacity.

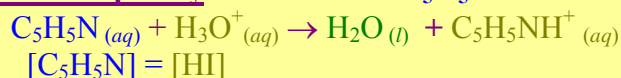
Example 8: Calculate the number of moles of HI_(aq) that must be added to 0.500 L of 2.00 mol/L of C₅H₅NHI_(aq) (K_b of C₅H₅N = 1.7×10^{-9}) to produce a solution buffered at pH = 4.75.



C₅H₅NH⁺ is a weak acid. Hence, we need to write the acid dissociation reaction and obtain K_a .



Adding HI (or H₃O⁺) will **react completely** with the base C₅H₅N:



For the buffered solution, we can use the Henderson-Hasselbalch Equation to find $\left(\frac{[\text{C}_5\text{H}_5\text{N}]}{[\text{C}_5\text{H}_5\text{NH}^+]} \right)$.

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{C}_5\text{H}_5\text{N}]}{[\text{C}_5\text{H}_5\text{NH}^+]} \right) \qquad \text{pH} - \text{p}K_a = \log \left(\frac{[\text{C}_5\text{H}_5\text{N}]}{[\text{C}_5\text{H}_5\text{NH}^+]} \right)$$

$$4.75 - [-\log(5.882 \times 10^{-6})] = \log \left(\frac{[\text{C}_5\text{H}_5\text{N}]}{2.00 \text{ M}} \right)$$

$$\log \left(\frac{[\text{C}_5\text{H}_5\text{N}]}{2.00 \text{ M}} \right) = -0.4804749798$$

$$10^{-0.4804749798} = \left(\frac{[\text{C}_5\text{H}_5\text{N}]}{2.00 \text{ M}} \right)$$

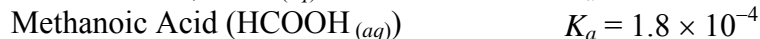
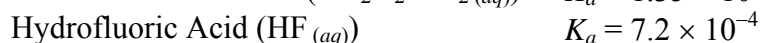
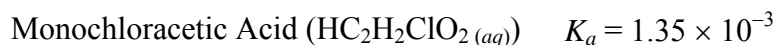
$$\text{Since } [\text{C}_5\text{H}_5\text{N}] = [\text{HI}] \qquad [\text{C}_5\text{H}_5\text{N}] = (2.00 \text{ M})(10^{-0.4804749798})$$

$[\text{HI}] = 0.662 \text{ M} \longleftarrow [\text{C}_5\text{H}_5\text{N}] = 0.662 \text{ M}$

$$n_{\text{HI}} = [\text{HI}]V = (0.662 \text{ mol/L})(0.500 \text{ L})$$

$$n_{\text{HI}} = 0.331 \text{ mol}$$

Example 9: From a list of the K_a below, select the buffered solution that has a pH of around 3.20. Write up a procedure that makes 1.00 L of this particular buffered solution.



For Optimal Buffering Capacity, $[\text{HA}] = [\text{A}^-]$, which means $\text{pH} = \text{p}K_a$.

$$\text{p}K_a = -\log K_a$$

$$K_a = 10^{-\text{p}K_a} = 10^{-(3.20)} \qquad K_a = 6.31 \times 10^{-4} \text{ (Closest Weak Acid to this } K_a \text{ is HF}_{(aq)})$$

Possible Procedure to Make 2.00 M of F⁻/HF Buffered Solution.

(Large concentrations are needed for good buffering capacity.)

1. Obtain 1.00 L of 2.00 M of HF. (May entail dilution from stock solution.)

2. Dissolve 83.98 g of NaF ($m = nM = 2.00 \text{ mol/L} \times 1.00 \text{ L} \times 41.99 \text{ g/mol}$) into the 1.00 L of 2.00 M HF_(aq)

Assignment

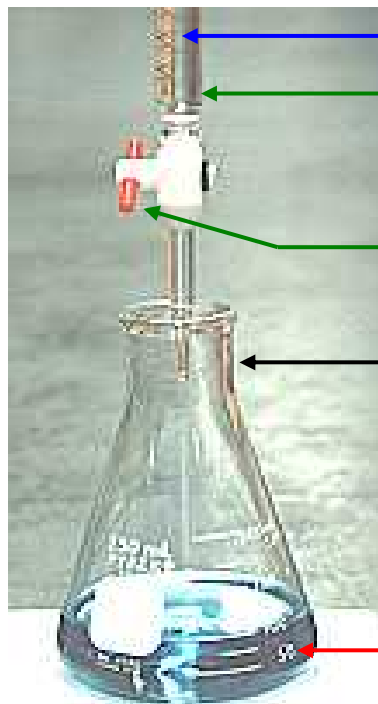
16.1 & 16.2 pg. 740 #1 to 6; #7 to 20; pg. 744 #102 and 107

16.4: Acid-Base Titrations

pH Curve: - a graph that shows the how the pH changes as the titration proceeds (as titrant is to the analyzed).

Titration: - a volumetric analysis that involves measuring the volume of known concentration solution to measure a given volume of an unknown concentration solution.

Titration Set-up



Titrant: - the solution of known concentration.

Buret: - a precise apparatus to deliver the titrant.
- the volume of the titrant added is read by subtracting the final volume and the initial volume.

Buret Valve: - can be adjusted to let one drop out at a time.

Erlenmeyer Flask: - a container commonly uses to hold the analyte. (Narrow mouth prevents splash and spillage.)

Analyte: - the solution of an unknown concentration.
- the exact volume is usually delivered by a pipet.

Acid-Base Titration: - volumetric analysis that assist in determining the unknown concentration in an acid and base neutralization.

Equivalent Point (Stoichiometric Point): - a point where the number of moles of H^+ is equivalent to the number of moles of OH^- . ($n_{H^+} = n_{OH^-}$)

Endpoint: - a point where the indicator actually changes colour to indicate neutralization is completed.

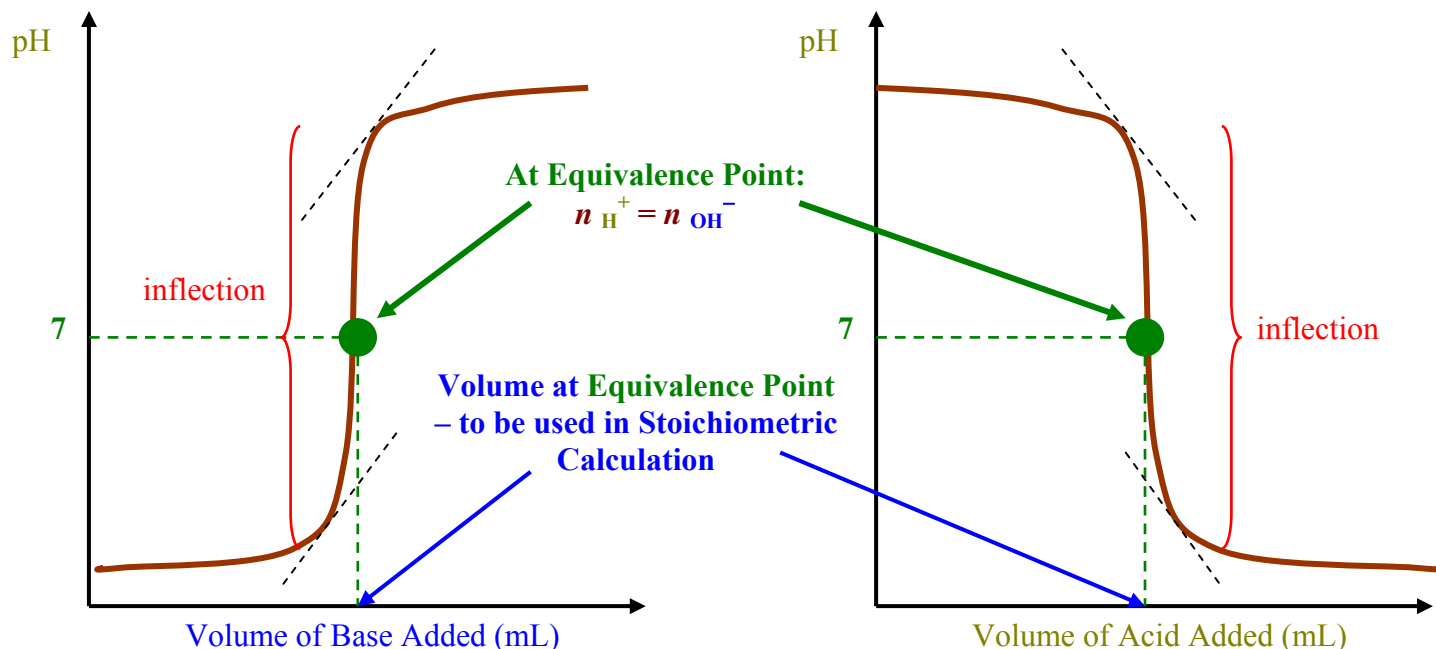
Indicator: - a chemical that changes colour due to the pH of the solution (more in the next section).

Inflection: - the part of the curve where there is a sudden rise or drop in pH.
- the midpoint of the inflection marks the equivalence point.

1. **Titration Between Strong Acids and Strong Base:** - Equivalence Point always occur at $\text{pH} = 7$.

Strong Acid (unknown concentration with known volume - analyte) titrated with Strong Base (known concentration - titrant)

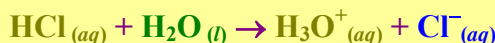
Strong Base (unknown concentration and known volume - analyte) titrated with Strong Acid (known concentration - titrant)



Example 1: Calculate the pH when 30.0 mL of 0.100 M of $\text{HCl}_{(aq)}$ is titrated with 0.200 M of $\text{KOH}_{(aq)}$ at:

- 0 mL of $\text{KOH}_{(aq)}$ added.
- 5.00 mL of $\text{KOH}_{(aq)}$ added.
- 15.00 mL of $\text{KOH}_{(aq)}$ added.
- 20.00 mL of $\text{KOH}_{(aq)}$ added.

a. At 0 mL of $\text{KOH}_{(aq)}$ added, the pH will be solely based on the $[\text{HCl}]$ in the flask:



$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log(0.100)$$

$$\text{pH} = 1.00$$

b. At 5.00 mL of $\text{KOH}_{(aq)}$ added: (Before Stoichiometric Point)

Chemical Species Present:

SA	A
H_3O^+ , Cl^- , K^+ , OH^- , H_2O	
B	SB

Net-Ionic Equation: $\text{H}_3\text{O}^+_{(aq)} + \text{OH}^-_{(aq)} \rightarrow 2 \text{H}_2\text{O}_{(l)}$

	n of H_3O^+	n of OH^-
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(5.00 mL) = 1.00 mmol (LR)
Change	- 1.00 mmol	- 1.00 mmol
After	2.00 mmol	0.00 mmol

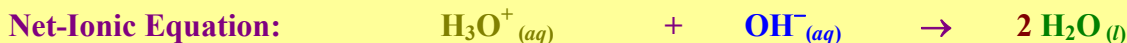
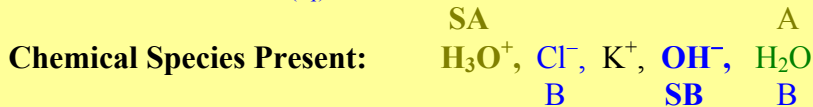
$$[\text{H}_3\text{O}^+] = \frac{n_{\text{H}_3\text{O}^+}}{\text{Total Volume}} = \frac{2.00 \text{ mmol}}{(30.00 \text{ mL} + 5.00 \text{ mL})} = 0.05714 \text{ mol/L}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log(0.05714)$$

$$\text{pH} = 1.24$$

(pH's results from similar Calculations with other volumes before the stoichiometric point is listed on the next page.)

c. At 15.00 mL of KOH_(aq) added: **(Stoichiometric Point)**



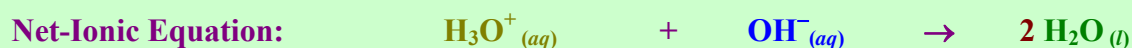
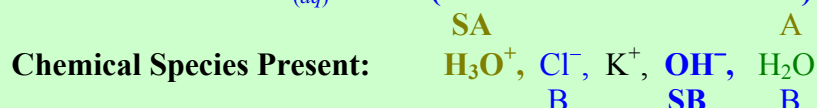
	n of H_3O^+	n of OH^-
Before	$(0.100 \text{ mol/L})(30.00 \text{ mL}) = 3.00 \text{ mmol}$	$(0.200 \text{ mol/L})(15.00 \text{ mL}) = 3.00 \text{ mmol}$
Change	$- 3.00 \text{ mmol}$	$- 3.00 \text{ mmol}$
After	0.00 mmol	0.00 mmol

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.00 \times 10^{-7} \text{ mol/L}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log(1.00 \times 10^{-7})$$

$$\text{pH} = 7.00$$

d. At 20.00 mL of KOH_(aq) added: **(After Stoichiometric Point)**



	n of H_3O^+	n of OH^-
Before	$(0.100 \text{ mol/L})(30.00 \text{ mL}) = 3.00 \text{ mmol}$ (Limiting Reagent)	$(0.200 \text{ mol/L})(20.00 \text{ mL}) = 4.00 \text{ mmol}$
Change	$- 3.00 \text{ mmol}$	$- 3.00 \text{ mmol}$
After	0.00 mmol	1.00 mmol

$$[\text{OH}^-] = \frac{n_{\text{OH}^-}}{\text{Total Volume}} = \frac{1.00 \text{ mmol}}{(30.00 \text{ mL} + 20.00 \text{ mL})} = 0.0200 \text{ mol/L}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log(0.0200)$$

$$\text{pOH} = 1.70$$

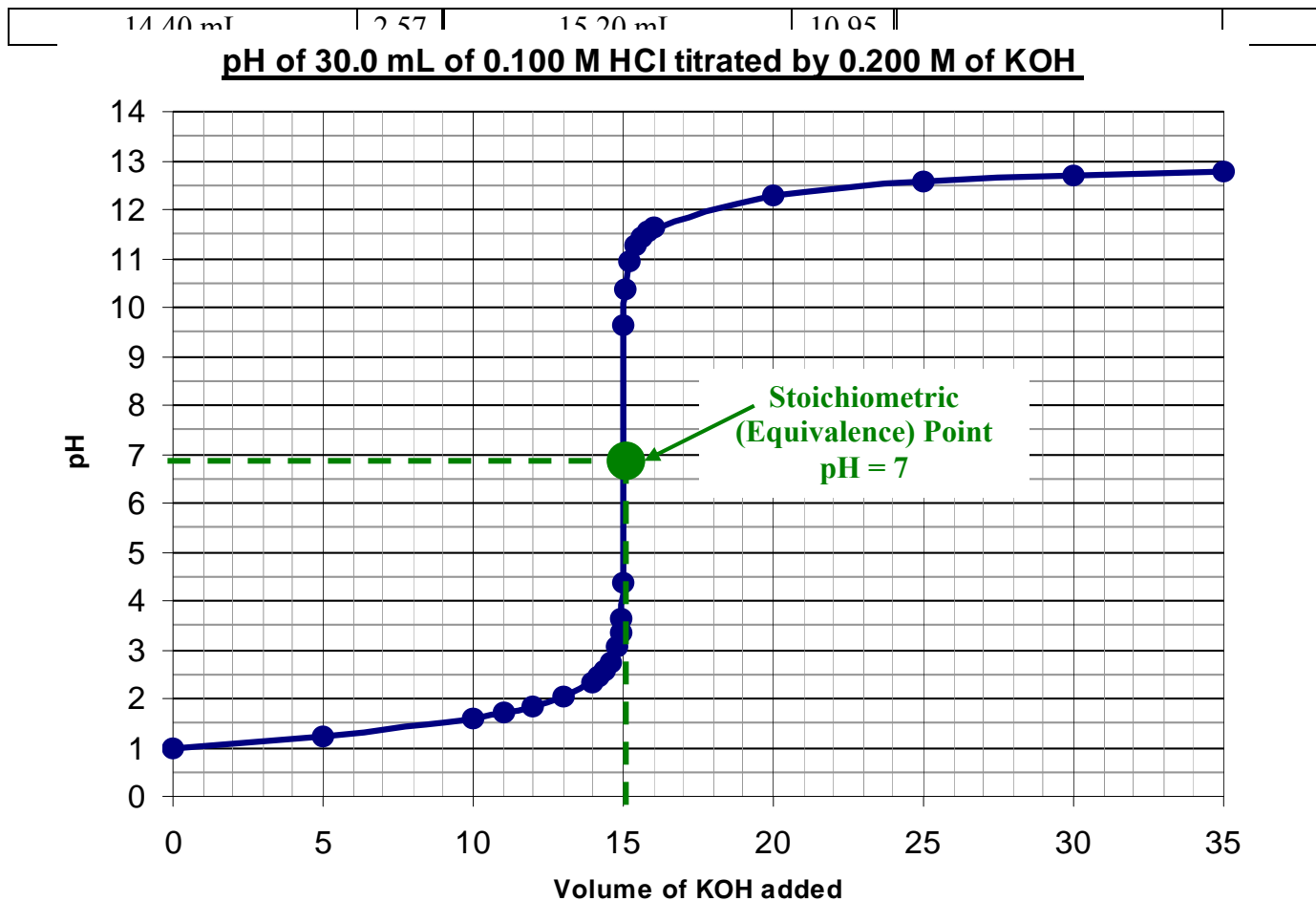
$$\text{pH} = 14 - \text{pOH}$$

(pH's results from similar Calculations with other volumes after the stoichiometric point is listed below.)

$$\text{pH} = 12.30$$

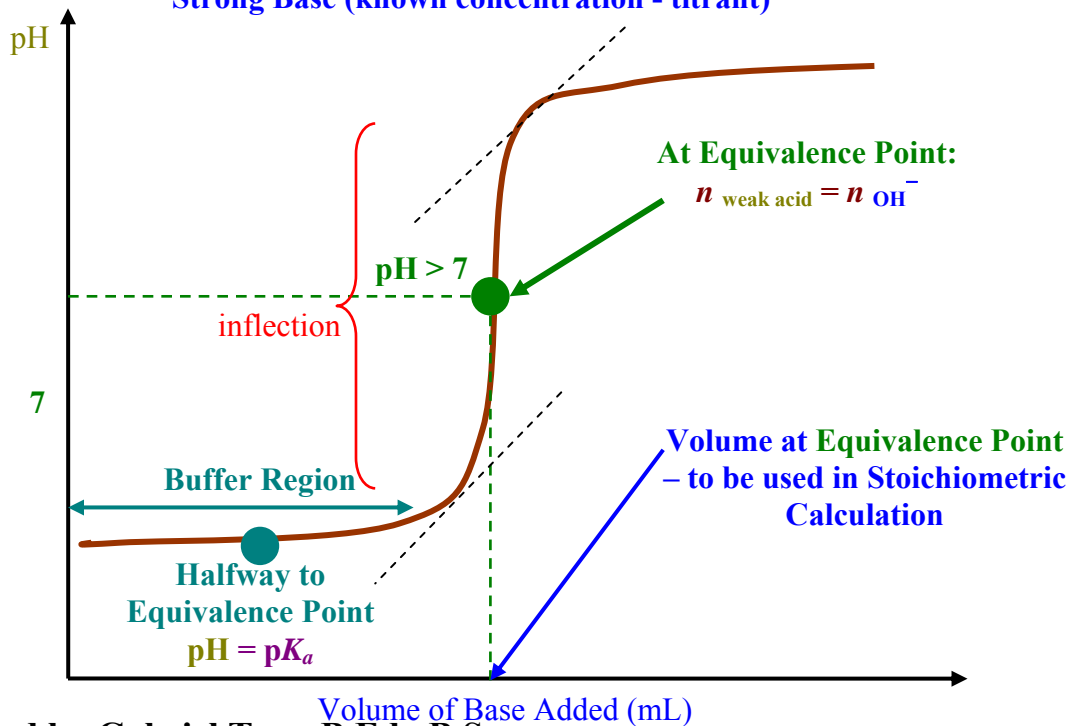
pH of 30.0 mL of 0.100 M HCl_(aq) titrated by 0.200 M of KOH_(aq)

Volume of KOH added	pH	Volume of KOH added	pH	Volume of KOH added	pH
0.00 mL	1.00	14.60 mL	2.75	15.40 mL	11.25
5.00 mL	1.24	14.80 mL	3.05	15.60 mL	11.42
10.00 mL	1.60	14.90 mL	3.35	15.80 mL	11.54
11.00 mL	1.71	14.95 mL	3.65	16.00 mL	11.64
12.00 mL	1.85	14.99 mL	4.35	20.00 mL	12.30
13.00 mL	2.03	15.00 mL (Equivalence)	7.00	25.00 mL	12.56
14.00 mL	2.34	15.01 mL	9.65	30.00 mL	12.70
14.20 mL	2.44	15.05 mL	10.35	35.00 mL	12.79



2. **Titration Between Weak Acids and Strong Base:** - Equivalence Point always occur at $pH > 7$.

Weak Acid (unknown concentration with known volume - analyte) titrated with Strong Base (known concentration - titrant)



c. At 15.00 mL of KOH (aq) added: (Stoichiometric Point)

Chemical Species Present: HCOOH, K⁺, OH⁻, H₂O
SA SB A/B

Net-Ionic Equation: HCOOH (aq) + OH⁻ (aq) → H₂O (l) + HCOO⁻ (aq)

	n of HCOOH	n of OH ⁻	n of HCOO ⁻
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(15.00 mL) = 3.00 mmol	0 mmol
Change	- 3.00 mmol	- 3.00 mmol	+ 3.00 mmol
After	0 mmol	0 mmol	3.00 mmol

$$[\text{HCOO}^-] = \frac{n_{\text{HCOO}^-}}{\text{Total Volume}} = \frac{3.00 \text{ mmol}}{(30.00 \text{ mL} + 15.00 \text{ mL})} = 0.0667 \text{ mol/L}$$

At Stoichiometric Point, after all the OH⁻ and HCOOH are used up, the resulting solution follows the regular weak base dissociation.

	HCOO ⁻ (aq)	H ₂ O (l)	HCOOH (aq)	OH ⁻ (aq)
Initial	0.0667 M		0	0 M
Change	-x		+x	+x
Equilibrium	(0.0667 - x)		x	x

CAN use Approximation:

$$\frac{[\text{HCOO}^-]_b}{K_b} = \frac{0.0667 \text{ M}}{5.556 \times 10^{-11}} = 1.2 \times 10^9 \geq 1000$$

Use 0.0667 in the denominator, because (0.0667 - x) ≈ 0.03667 [x is so small compared to 0.0667 M].

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-4}} \quad K_b = 5.556 \times 10^{-11} = \frac{[\text{HCOOH}][\text{OH}^-]}{[\text{HCOO}^-]}$$

$$5.556 \times 10^{-11} = \frac{(x)(x)}{(0.0667 - x)} \approx \frac{x^2}{0.0667}$$

$$[\text{OH}^-] = x = 1.92 \times 10^{-6} \text{ mol/L}$$

$$\text{pOH} = -\log(1.92 \times 10^{-6}) = 5.72$$

$$\text{pH} = 14 - \text{pOH} = 14 - 5.72$$

$$\text{pH} = 8.28$$

d. At 20.00 mL of KOH (aq) added: (After Stoichiometric Point)

Chemical Species Present: HCOOH, K⁺, OH⁻, H₂O
SA SB A/B

Net-Ionic Equation: HCOOH (aq) + OH⁻ (aq) → H₂O (l) + HCOO⁻ (aq)

	n of HCOOH	n of OH ⁻
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol (LR)	(0.200 mol/L)(20.00 mL) = 4.00 mmol
Change	- 3.00 mmol	- 3.00 mmol
After	0 mmol	1.00 mmol

After all the HCOOH is used up, the resulting solution has the following major species.

Chemical Species Present: K⁺, OH⁻, H₂O
SB A/B

$$[\text{OH}^-] = \frac{n_{\text{OH}^-}}{\text{Total Volume}} = \frac{1.00 \text{ mmol}}{(30.00 \text{ mL} + 20.00 \text{ mL})} = 0.0200 \text{ mol/L} \quad \text{pOH} = -\log(0.0200) = 1.70$$

$$\text{pH} = 14 - \text{pOH} = 14 - 1.70$$

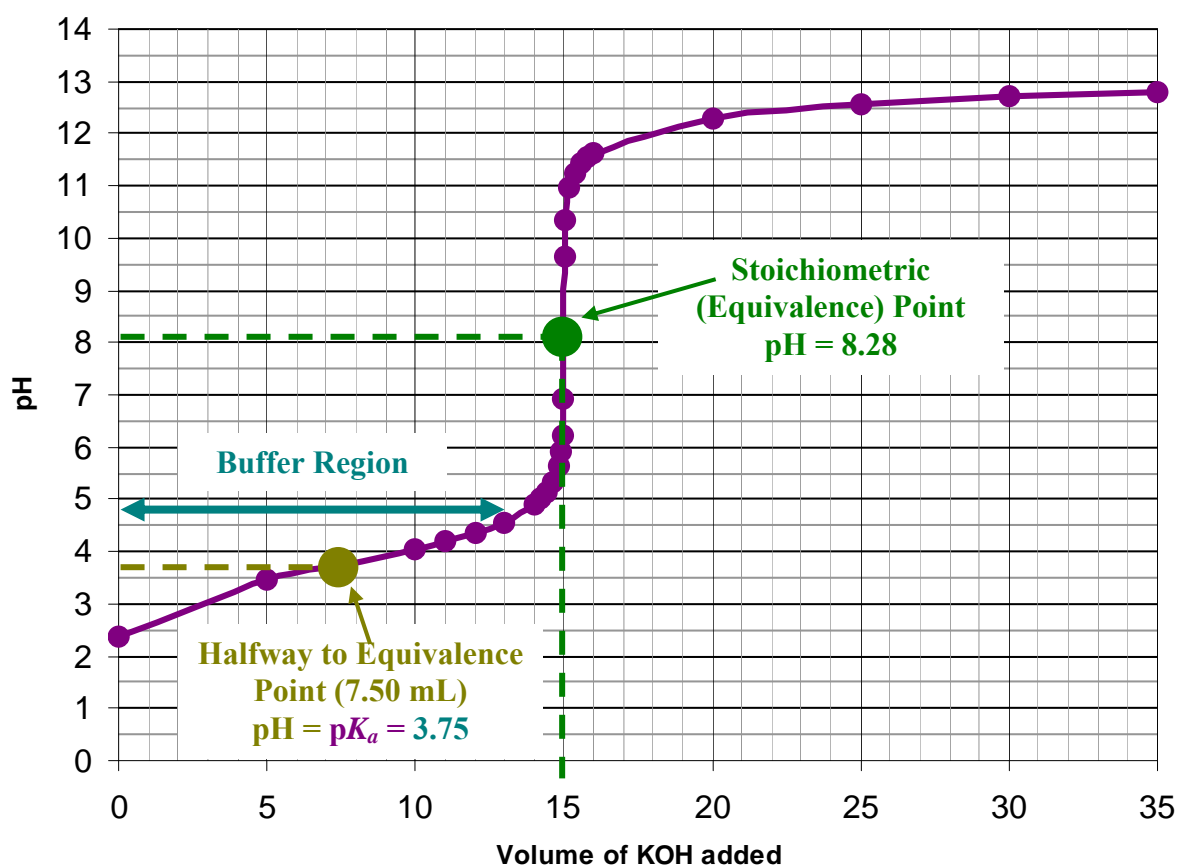
$$\text{pH} = 12.30$$

(pH's results from similar Calculations with other volumes after the stoichiometric point is listed on the next page.)

pH of 30.0 mL of 0.100 M HCOOH_(aq) titrated by 0.200 M of KOH_(aq)

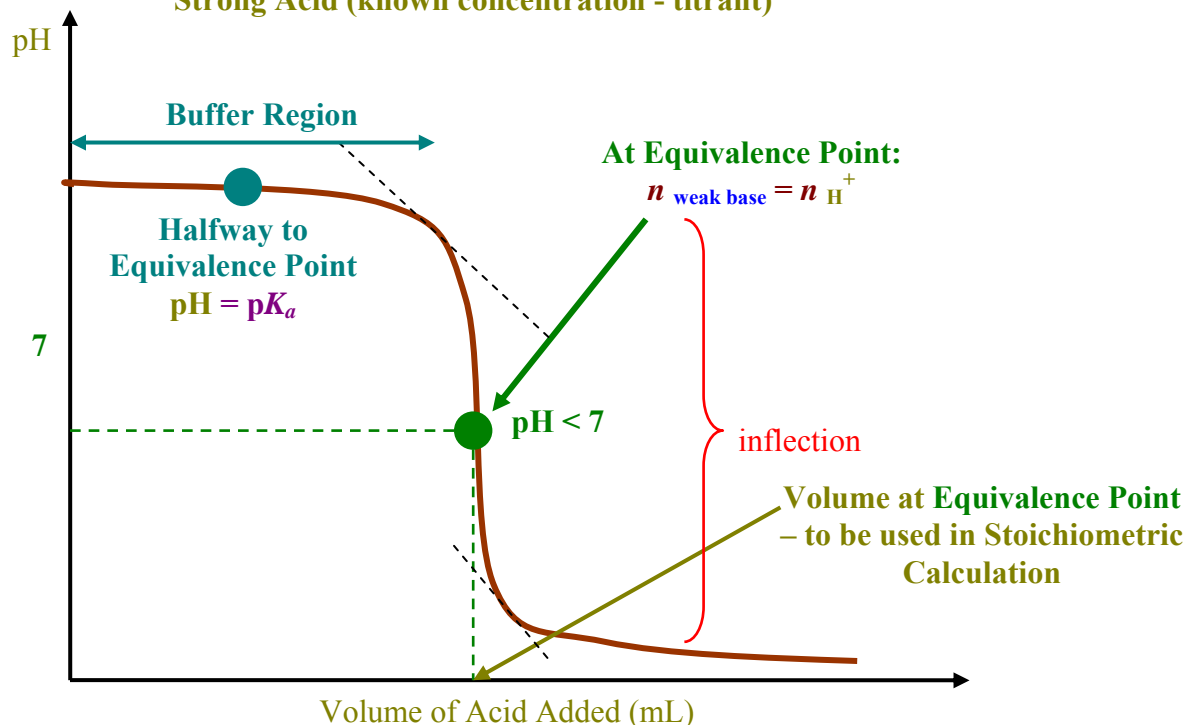
Volume of KOH added	pH	Volume of KOH added	pH	Volume of KOH added	pH
0.00 mL	2.38	14.40 mL	5.13	15.20 mL	10.95
5.00 mL	3.45	14.60 mL	5.31	15.40 mL	11.25
7.50 mL (Halfway to Equivalence)	3.75	14.80 mL	5.62	15.60 mL	11.42
10.00 mL	4.05	14.90 mL	5.91	15.80 mL	11.54
11.00 mL	4.19	14.95 mL	6.22	16.00 mL	11.64
12.00 mL	4.35	14.99 mL	6.92	20.00 mL	12.30
13.00 mL	4.56	15.00 mL (Equivalence)	8.28	25.00 mL	12.56
14.00 mL	4.89	15.01 mL	9.65	30.00 mL	12.70
14.20 mL	5.00	15.05 mL	10.35	35.00 mL	12.79

pH of 30.0 mL of 0.100 M HCOOH titrated by 0.200 M of KOH



3. Titration Between Weak Bases and Strong Acid: - Equivalence Point always occur at $\text{pH} < 7$.

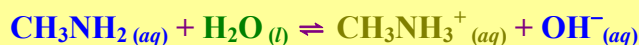
Weak Base (unknown concentration with known volume - analyte) titrated with Strong Acid (known concentration - titrant)



Example 3: Calculate the pH when 30.0 mL of 0.100 M of $\text{CH}_3\text{NH}_2(aq)$ ($K_b = 4.38 \times 10^{-4}$) is titrated with 0.200 mol/L of $\text{HCl}(aq)$ at:

- 0 mL of $\text{HCl}(aq)$ added.
- 7.50 mL of $\text{HCl}(aq)$ added.
- 15.00 mL of $\text{HCl}(aq)$ added.
- 20.00 mL of $\text{HCl}(aq)$ added.

a. At 0 mL of $\text{HCl}(aq)$ added, the pH will be calculated the weak base dissociation in the flask:



	$[\text{CH}_3\text{NH}_2]$	$[\text{CH}_3\text{NH}_3^+]$	$[\text{OH}^-]$
Initial	0.100 M	0	0 M
Change	-x	+x	+x
Equilibrium	(0.1 - x)	x	x

CANNOT use Approximation:

$$\frac{[\text{CH}_3\text{NH}_2]_0}{K_b} = \frac{0.100 \text{ M}}{4.38 \times 10^{-4}} = 228.3 < 1000$$

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

$$4.38 \times 10^{-4} = \frac{(x)(x)}{(0.1-x)} = \frac{x^2}{(0.1-x)}$$

$$[\text{OH}^-] = x = 0.00640 \text{ mol/L}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log(0.00640) = 2.19$$

$$\text{pH} = 14 - \text{pOH} = 14 - 2.19$$

```
solve(X^2/(0.1-X)
-4.38E-4,X,0,0,
0.1)
.0064027793
```

$$\text{pH} = 11.81$$

b. At 7.50 mL of $\text{HCl}_{(aq)}$ added: (Before Stoichiometric Point: at Halfway to Equivalence Point)

Chemical Species Present: CH_3NH_2 , H_3O^+ , Cl^- , H_2O
 SB SA A/B

Net-Ionic Equation: $\text{CH}_3\text{NH}_2_{(aq)} + \text{H}_3\text{O}^+_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)} + \text{CH}_3\text{NH}_3^+_{(aq)}$

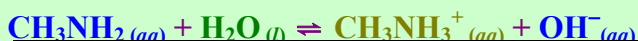
	n of CH_3NH_2	n of H_3O^+	n of CH_3NH_3^+
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(7.50 mL) = 1.50 mmol (LR)	0 mmol
Change	- 1.50 mmol	- 1.50 mmol	+ 1.50 mmol
After	1.50 mmol	0 mmol	1.50 mmol

$$[\text{CH}_3\text{NH}_2] = \frac{n_{\text{CH}_3\text{NH}_2}}{\text{Total Volume}} = \frac{1.50 \text{ mmol}}{(30.00 \text{ mL} + 7.50 \text{ mL})} = 0.0400 \text{ mol/L}$$

$$[\text{CH}_3\text{NH}_3^+] = \frac{n_{\text{CH}_3\text{NH}_3^+}}{\text{Total Volume}} = \frac{1.50 \text{ mmol}}{(30.00 \text{ mL} + 7.50 \text{ mL})} = 0.0400 \text{ mol/L}$$

} $[\text{CH}_3\text{NH}_3^+] = [\text{CH}_3\text{NH}_2]$
 $\text{pH} \approx \text{p}K_a$
 (Max Buffer Capacity)

After all the H_3O^+ is used up, the resulting solution follows the regular weak base dissociation.



	$[\text{CH}_3\text{NH}_2]$	$[\text{CH}_3\text{NH}_3^+]$	$[\text{OH}^-]$
Initial	0.0400 M	0	0.0400 M
Change	-x	+x	+x
Equilibrium	(0.04 - x)	x	(0.04 + x)

CANNOT use Approximation:

$$\frac{[\text{CH}_3\text{NH}_2]_0}{K_b} = \frac{0.0400 \text{ M}}{4.38 \times 10^{-4}} = 91.3 < 1000$$

$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} \quad 4.38 \times 10^{-4} = \frac{(x)(0.04 + x)}{(0.04 - x)}$$

$$[\text{OH}^-] = x = 4.29 \times 10^{-4} \text{ mol/L}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log(4.29 \times 10^{-4}) = 3.37$$

$$\text{pH} = 14 - \text{pOH} = 14 - 3.37$$

(pH's results from similar Calculations with other volumes before the stoichiometric point is listed on the next page.)

$\text{pH} = 10.63$

c. At 15.00 mL of $\text{HCl}_{(aq)}$ added: (Stoichiometric Point)

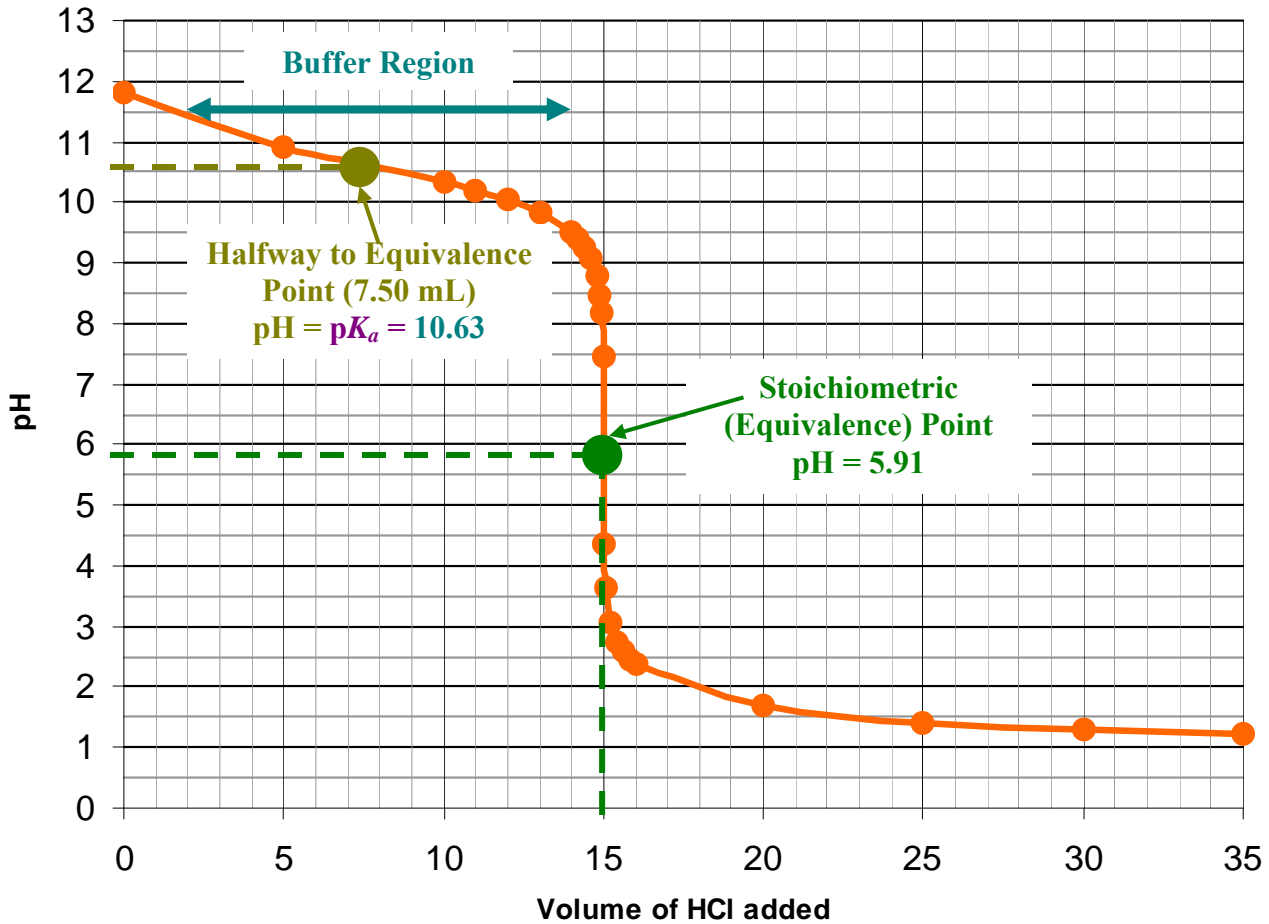
Chemical Species Present: CH_3NH_2 , H_3O^+ , Cl^- , H_2O
 SB SA A/B

Net-Ionic Equation: $\text{CH}_3\text{NH}_2_{(aq)} + \text{H}_3\text{O}^+_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)} + \text{CH}_3\text{NH}_3^+_{(aq)}$

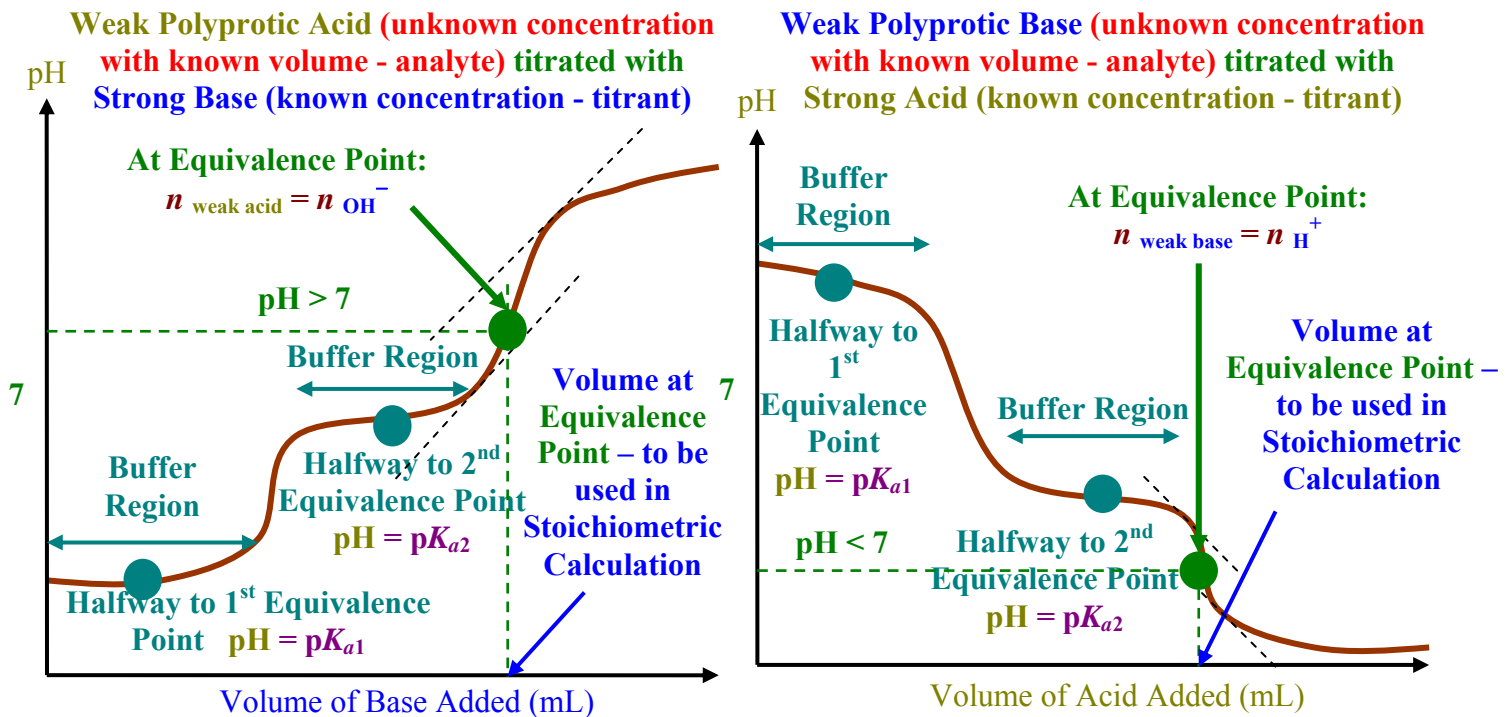
	n of CH_3NH_2	n of H_3O^+	n of CH_3NH_3^+
Before	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(15.00 mL) = 3.00 mmol	0 mmol
Change	- 3.00 mmol	- 3.00 mmol	+ 3.00 mmol
After	0 mmol	0 mmol	3.00 mmol

$$[\text{CH}_3\text{NH}_3^+] = \frac{n_{\text{CH}_3\text{NH}_3^+}}{\text{Total Volume}} = \frac{3.00 \text{ mmol}}{(30.00 \text{ mL} + 15.00 \text{ mL})} = 0.0667 \text{ mol/L}$$

pH of 30.0 mL of 0.100 M Methylamine titrated by 0.200 M of HCl

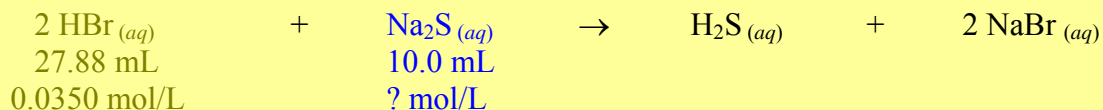


4. **Titration Between Strong Acid/Bases and Polyprotic Acid/Base:** - there are multiple equivalence points.
 - the last equivalence point indicates the stoichiometric volume.



Example 4: Calculate the concentration of 10.00 mL of $\text{Na}_2\text{S}_{(aq)}$ is titrated with 0.0350 M of $\text{HBr}_{(aq)}$ when the first and second equivalence points are at 13.34 mL and 27.88 mL respectively.

For Acid-Base Stoichiometry, do NOT write the Net-Ionic Equation. Write the Molecular Equation to do Stoichiometry. We use the second equivalence point because Na_2S can accept two protons.



$$\textcircled{1} n_{\text{HBr}} = CV = (0.0350 \text{ mol/L}) (27.88 \text{ mL}) = 0.9758 \text{ mmol}$$

$$\textcircled{2} n_{\text{Na}_2\text{S}} = 0.9758 \text{ mmol HBr} \times \frac{1 \text{ mol Na}_2\text{S}}{2 \text{ mol HBr}} = 0.4879 \text{ mmol Na}_2\text{S}$$

$$\textcircled{3} [\text{Na}_2\text{S}] = \frac{n}{V} = \frac{0.4879 \text{ mmol}}{10.0 \text{ mL}} = 0.04879 \text{ mol/L}$$

$$[\text{Na}_2\text{S}] = 0.0488 \text{ mol/L}$$

16.5: Acid-Base Indicators

Acid-Base Indicators: - chemicals that change colours at a specific pH range.

- they are themselves organic acids. Since they are usually very big structurally, we usually use abbreviations to describe them in chemical equations.
- this is due to the acidic form of the indicator (HIn) has a different colour than its basic form (In^-).

- the colour change occurs when $\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{1}{10}$ (titrating an acid with a base) or when

$$\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{10}{1} \text{ (titrating a base with an acid).}$$

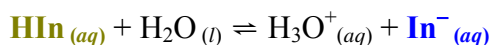
Using the Henderson-Hasselbach Equation to determine pH Range for Indicators to Change Colour:

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{In}^-]}{[\text{HIn}]}\right) \quad \text{(Substitute observable Colour Change Ratio } \frac{[\text{In}^-]}{[\text{HIn}]} = \frac{1}{10} \text{ or } \frac{10}{1}\text{)}$$

$$\text{pH} = \text{p}K_a + \log\left(\frac{1}{10}\right) \quad \text{or} \quad \text{pH} = \text{p}K_a + \log\left(\frac{10}{1}\right)$$

$$\text{pH} = \text{p}K_a - 1 \quad \text{or} \quad \text{pH} = \text{p}K_a + 1$$

pH Range for Indicators to Change Colour



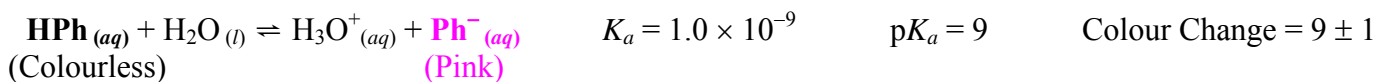
$$\text{pH} = \text{p}K_a \pm 1$$

$$\text{pH} \leq (\text{p}K_a - 1) \rightarrow \text{Colour of HIn}_{(aq)}$$

$$\text{pH} \geq (\text{p}K_a + 1) \rightarrow \text{Colour of In}^-_{(aq)}$$

$$\text{pH} = \text{p}K_a \rightarrow \text{Mixed Colours of HIn}_{(aq)} + \text{In}^-_{(aq)}$$

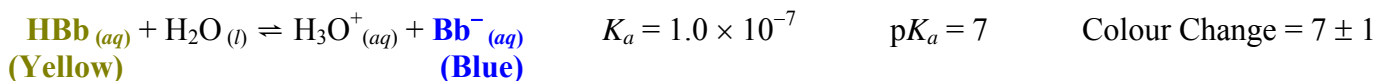
Example: Phenolphthalein ($\text{HPh}_{(aq)}/\text{Ph}^{-}_{(aq)}$) changes colours from colourless to pink at a pH range of 8.0 to 10.0. At its **mid-range (9.0 – average of 8.0 and 10.0)**, the colour would be **light pink (colourless + pink)**.



At $\text{pH} \leq 9$, phenolphthalein is colourless.

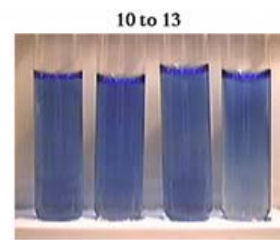
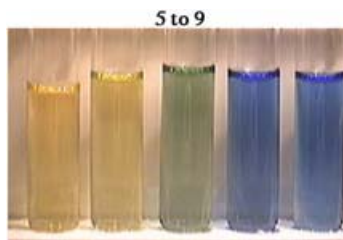
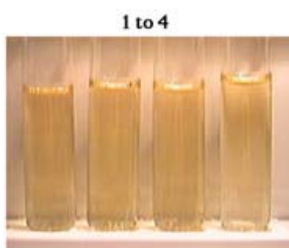
At $\text{pH} \geq 10$, phenolphthalein is pink.

Example: Bromothymol Blue ($\text{HBb}_{(aq)}/\text{Bb}^{-}_{(aq)}$) changes colours from yellow to blue at a pH range of 6.0 to 8.0. At its **mid-range (7.0 – average of 6.0 and 8.0)**, the colour would be **green (yellow + blue)**.



At $\text{pH} \leq 6$, bromothymol blue is yellow.

At $\text{pH} \geq 8$, bromothymol blue is blue.



Example 1: The pH curve of methylamine with HCl has an equivalence point at $\text{pH} = 5.91$. Using the following information below, decide on an indicator to use in this titration and state what colour change the experimenter should be looking for.

Indicators	Colour Change	K_a
methyl orange	red to yellow	1.6×10^{-4}
methyl red	red to yellow	1.0×10^{-5}
bromocresol purple	yellow to purple	2.0×10^{-6}
bromothymol blue	yellow to blue	1.0×10^{-7}

First, we need to decide on the pH range that these indicators will change colors.

Indicators	Colour Change	K_a	pH range ($\text{p}K_a \pm 1$)
methyl orange	red to yellow	1.6×10^{-4}	3.8 ± 1 (2.8 to 4.8)
methyl red	red to yellow	1.0×10^{-5}	5.0 ± 1 (4.0 to 6.0)
bromocresol purple	yellow to purple	2.0×10^{-6}	5.7 ± 1 (4.7 to 6.7)
bromothymol blue	yellow to blue	1.0×10^{-7}	7.0 ± 1 (6.0 to 8.0)

We can see that $\text{pH} = 5.91$ is well within the color change of **bromocresol purple** (for methyl red, it is too close to the boundary point). **The color where the experimenter will stop is brown (yellow + purple) with the $\text{pH} = 5.7$ (end-point).**

Assignment

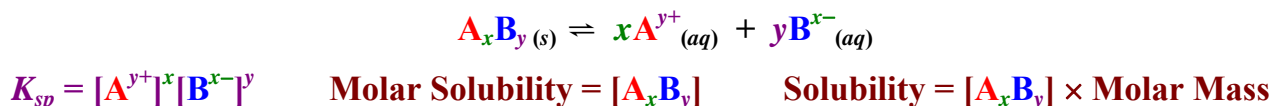
16.4 pg. 741 #22 to 32; pg. 734–744 #88 and 106
16.5 pg. 741 #33 to 38; pg. 744 #100

16.6: Solubility Equilibria

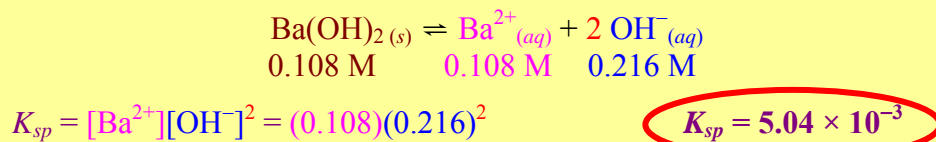
Solubility Product (K_{sp}): - the equilibrium constant as a salt dissolves into its aqueous ions.
 - sometimes refer to as **solubility product constant**.
 - like equilibrium constant, **K_{sp} is unitless**.

Molar Solubility: - the equilibrium position of the solvation equilibrium.
 - the maximum amount of salts in moles dissolved per Litre of solvent.

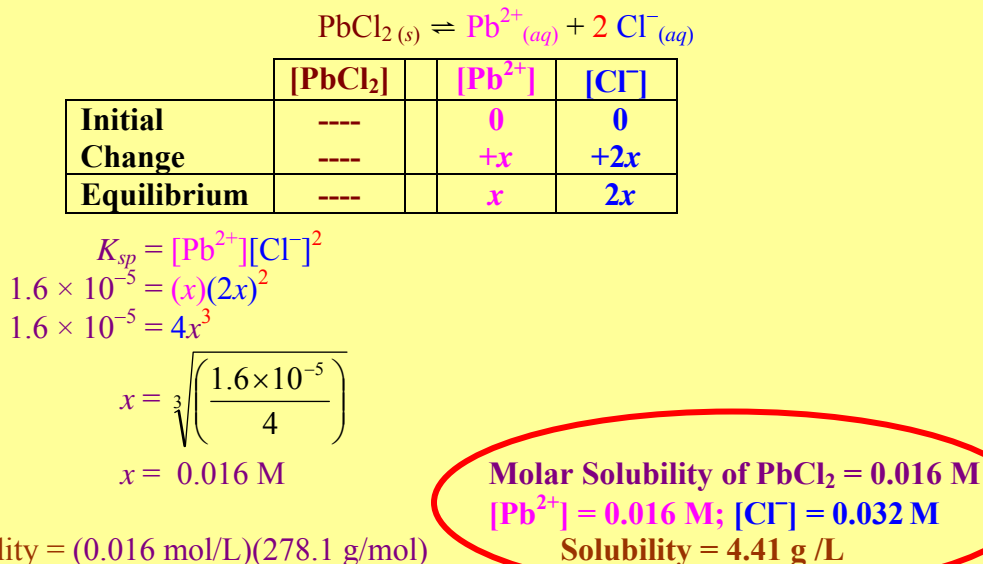
Solubility: - the maximum mass of salts in moles dissolved per Litre of solvent.



Example 1: The molar solubility of $Ba(OH)_2 (s)$ is 0.108 mol/L. Determine the molar concentration of each ion and the solubility product of $Ba(OH)_2 (s)$.



Example 2: The K_{sp} for lead (II) chloride is 1.6×10^{-5} . Calculate the molar concentrations of each ion and the molar solubility and the solubility of solid lead (II) chloride.

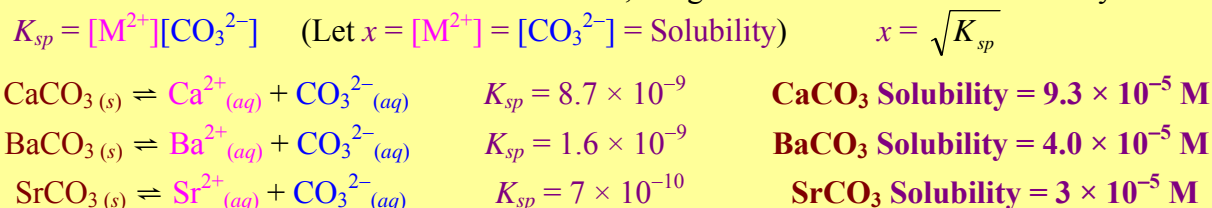


Relative Solubility: - how the solubilities of various salts compare.

- **the solubilities of salts can only be compared directly from K_{sp} values when the salts produce the same number of ions.** (In general, the bigger the K_{sp} , the higher the solubility.)
- **the solubilities of salts that produce Different Number of Ions CANNOT be compared directly from K_{sp} .** They must be calculated individually before comparison can be made.

Example 3: Compare the solubilities of CaCO_3 , BaCO_3 and SrCO_3 with the K_{sp} values as 8.7×10^{-9} , 1.6×10^{-9} , and 7×10^{-10} respectively.

Since all salts listed here have the same number of ions, the general solution for the solubility is as follows:



The Solubility increases from $\text{SrCO}_3 < \text{BaCO}_3 < \text{CaCO}_3$. (Since CaCO_3 has the largest K_{sp} and all salts have the same number of ions dissociated, CaCO_3 has the highest solubility.)

Example 4: Compare the solubilities of AgCl , Ag_2CO_3 , and Ag_3PO_4 with the K_{sp} values as 1.6×10^{-10} , 8.1×10^{-12} , 1.8×10^{-18} respectively.

Since all salts listed here have different number of ions, the solubility of each salt has to be calculated separately.



	[AgCl]	[Ag ⁺]	[Cl ⁻]
Initial	----	0	0
Change	----	+x	+x
Equilibrium	----	x	x

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

$$1.6 \times 10^{-10} = (x)(x)$$

$$1.6 \times 10^{-10} = x^2$$

$$x = \sqrt{1.6 \times 10^{-10}}$$

$$x = 1.3 \times 10^{-5} \text{ M}$$

Solubility of $\text{AgCl} = 1.3 \times 10^{-5} \text{ M}$



	[Ag ₂ CO ₃]	[Ag ⁺]	[CO ₃ ²⁻]
Initial	----	0	0
Change	----	+2x	+x
Equilibrium	----	2x	x

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

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

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

$$x = \sqrt[3]{\left(\frac{8.1 \times 10^{-12}}{4}\right)} = 1.3 \times 10^{-4} \text{ M}$$

Solubility of $\text{Ag}_2\text{CO}_3 = 1.3 \times 10^{-4} \text{ M}$



	[Ag ₃ PO ₄]	[Ag ⁺]	[PO ₄ ³⁻]
Initial	----	0	0
Change	----	+3x	+x
Equilibrium	----	3x	x

$$K_{sp} = [\text{Ag}^+]^3[\text{PO}_4^{3-}]$$

$$1.8 \times 10^{-18} = (3x)^3(x)$$

$$1.8 \times 10^{-18} = 27x^4$$

$$x = \sqrt[4]{\left(\frac{1.8 \times 10^{-18}}{27}\right)} = 1.6 \times 10^{-5} \text{ M}$$

Solubility of $\text{Ag}_3\text{PO}_4 = 1.6 \times 10^{-5} \text{ M}$

The Solubility increases from $\text{AgCl} < \text{Ag}_3\text{PO}_4 < \text{Ag}_2\text{CO}_3$. Again we can see here because these salts do not dissociate the same number of ions, we cannot compare their solubilities directly from their K_{sp} values.

16.7: Separation of Ions by Fractional Precipitation

Ion Product (Q): - similar to reaction quotient, it measures the initial ion concentrations and compares it with K_{sp} .

$$Q = [A^{y+}]_0^x [B^{x-}]_0^y$$

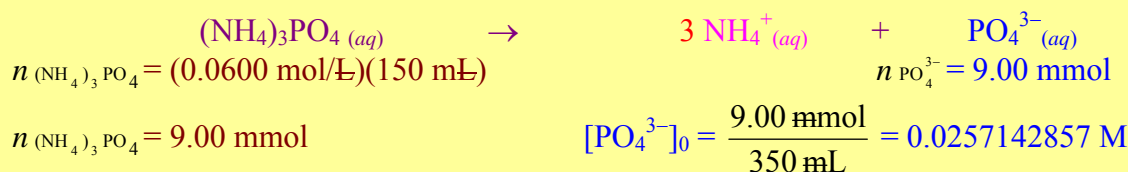
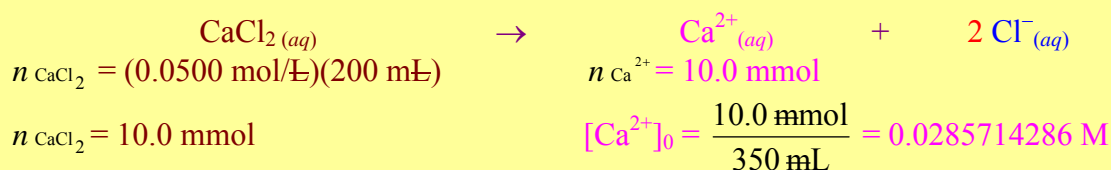
- a. **When $Q > K_{sp}$, the system will shift to the solid salt (reactant).** In this condition, Q indicates that **there are too much ions (products)**. Therefore, the system has to **Shift back to the Left and Precipitation occurs**.
- b. **When $Q < K_{sp}$, the system will shift to the ions (product).** In this condition, Q indicates that **there are too little ions (products)**. Therefore, the system has to **Shift to the Right and there will be NO Precipitation**.

Calculations involving Ion Product and Final Ion Concentrations:

- Determine the **Initial Concentrations** of the ions that will likely form a precipitate. (Remember to divide the moles by the total volume.)
- Calculate the **Ion product**.
- Compare it to the K_{sp} value** of the solid and decide whether precipitation will form.
- If precipitation occurs ($Q > K_{sp}$), then **write the net ionic equation for the precipitation**.
- Using regular stoichiometry**, run the reaction to completion and **determine the concentration of the excess ion**.
- Redo the equilibrium for dissolving**. Using K_{sp} and the excess ion concentration as initial concentration to find the **final concentrations of both ions**.

Example 1: 200 mL of 0.0500 M of calcium chloride is reacted with 150 mL of 0.0600 M of ammonium phosphate. Determine whether $\text{Ca}_3(\text{PO}_4)_2$ will precipitate and calculate the concentrations of Ca^{2+} and PO_4^{3-} in the final solution. The K_{sp} value of $\text{Ca}_3(\text{PO}_4)_2$ is 1.3×10^{-32} .

Both CaCl_2 and $(\text{NH}_4)_3\text{PO}_4$ dissociates completely in water.



Calculating Ion Product: $Q = [\text{Ca}^{2+}]_0^3 [\text{PO}_4^{3-}]_0^2 = (0.0285714286)^3 (0.0257142857)^2$
 $Q = 1.5 \times 10^{-8}$

$Q (1.5 \times 10^{-8}) > K_{sp} (1.3 \times 10^{-32})$ Precipitation Occurs

Running the precipitation reaction to completion and using regular stoichiometry,

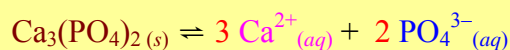


	n of Ca^{2+}	n of PO_4^{3-}
Before	10.0 mmol (LR)	9.00 mmol
Change	-10.00 mmol	$-\left(\frac{2}{3}\right)(10.0 \text{ mmol})$
After	0 mmol	2.33333333 mmol

$$\text{New } [\text{PO}_4^{3-}]_0 = \frac{2.33333333 \text{ mmol}}{350 \text{ mL}}$$

$$\text{New } [\text{PO}_4^{3-}]_0 = 0.019047619 \text{ M}$$

Finally, using K_{sp} and the new $[\text{PO}_4^{3-}]_0$,



	$[\text{Ca}_3(\text{PO}_4)_2]$	$[\text{Ca}^{2+}]$	$[\text{PO}_4^{3-}]$
Initial	----	0	0.019047619 M
Change	----	+3x	+2x
Equilibrium	----	3x	0.019047619 + 2x

$$K_{sp} = [\text{Ca}^{2+}]^3 [\text{PO}_4^{3-}]^2$$

$$1.3 \times 10^{-32} = (3x)^3 (0.019047619 + 2x)^2$$

$$1.3 \times 10^{-32} \approx (3x)^3 (0.019047619)^2$$

$$x^3 \approx \frac{1.3 \times 10^{-32}}{(3)^3 (0.019047619)^2}$$

$$x \approx \sqrt[3]{\frac{1.3 \times 10^{-32}}{(3)^3 (0.019047619)^2}} \approx 1.0989 \times 10^{-10}$$

$$[\text{Ca}^{2+}] = 3(1.0989 \times 10^{-10} \text{ M})$$

$$[\text{Ca}^{2+}] = 3.3 \times 10^{-10} \text{ M}$$

$$[\text{PO}_4^{3-}] = 0.019047619 \text{ M} + 2(1.0989 \times 10^{-10} \text{ M})$$

$$[\text{PO}_4^{3-}] = 0.019 \text{ M}$$

CAN use Approximation:

$$\frac{[\text{PO}_4^{3-}]_0^2}{K_{sp}} = \frac{(0.019047619 \text{ M})^2}{1.3 \times 10^{-32}}$$

$$= 2.8 \times 10^{28} \gg 1000$$

Because in the K_{sp} expression, PO_4^{3-} is squared, the rule of thumb must reflect this operation.

Use 0.019047619 in the denominator, because $(0.019047619 + 2x) \approx 0.019047619$ [$2x$ is so small compared to 0.019047619 M].

Assignment

16.6 pg. 741–744 #39 to 54, 92, 94, 99, 109, 112, 113, 116

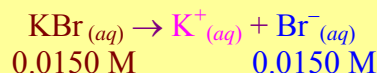
16.7 pg. 742–744 #55, 56, 93, 110

16.8: Common Ion Effect and Solubility

Common-Ion Effect and Solubility: - when a solution containing a common ion of a salt is used as a solvent, the **solubility of the salt will lower** as a result.
 - set up equilibrium ICE Box to assist in calculation.

Example 1: A 0.0150 M of $\text{KBr}_{(aq)}$ is used as a solvent for solid $\text{PbBr}_{2(s)}$. Calculate the final concentrations for both ions and the molar solubility of $\text{PbBr}_{2(s)}$ given that K_{sp} for PbBr_2 is 4.6×10^{-6} .

KBr dissociates completely in water.



	$\text{PbBr}_{2(s)}$	\rightleftharpoons	$\text{Pb}^{2+}_{(aq)}$	$+ 2 \text{Br}^-_{(aq)}$
	[PbBr ₂]		[Pb ²⁺]	[Br ⁻]
Initial	----		0	0.0150 M
Change	----		+x	+2x
Equilibrium	----		x	0.0150 + 2x

CANNOT use Approximation:

$$\frac{[\text{Br}^-]_0^2}{K_{sp}} = \frac{(0.0150 \text{ M})^2}{4.6 \times 10^{-6}} = 48.9 < 1000$$

Because in the K_{sp} expression, Br^- is squared, the rule of thumb must reflect this operation.

$$K_{sp} = [\text{Pb}^{2+}][\text{Br}^-]^2$$

$$4.6 \times 10^{-6} = (x)(0.0150 + 2x)^2$$

$$x = 0.0062 \text{ M}$$

```
solve(X(0.0150+2
X)^2-4.6E-6,X,0,(
0,.0150/2))
.006161609
```

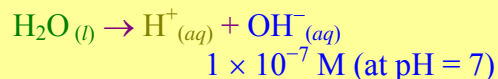
$$[\text{Br}^-] = 0.0150 \text{ M} + 2(0.0062 \text{ M})$$

Solubility of PbBr_2 in 0.0150 M of $\text{KBr}_{(aq)}$ = 0.0062 M
 $[\text{Pb}^{2+}] = 0.0062 \text{ M}$; $[\text{Br}^-] = 0.027 \text{ M}$

Note: If there was no common ion, the solubility of PbBr_2 would be 0.010 M. Hence, **with the presence of a common ion, the solubility of a salt is lowered.**

Example 2: Determine the final concentrations for both ions of $\text{Cu}(\text{OH})_{2(s)}$ in neutral water given that K_{sp} for $\text{Cu}(\text{OH})_2$ is 2.2×10^{-20} .

We have to consider the $[\text{OH}^-]$ in water.



	$\text{Cu}(\text{OH})_{2(s)}$	\rightleftharpoons	$\text{Cu}^{2+}_{(aq)}$	$+ 2 \text{OH}^-_{(aq)}$
	[Cu(OH) ₂]		[Pb ²⁺]	[OH ⁻]
Initial	----		0	$1 \times 10^{-7} \text{ M}$
Change	----		+x	+2x
Equilibrium	----		x	$(1 \times 10^{-7}) + 2x$

CAN use Approximation:

$$\frac{[\text{OH}^-]_0^2}{K_{sp}} = \frac{(1 \times 10^{-7} \text{ M})^2}{2.2 \times 10^{-20}} = 4.5 \times 10^5 < 1000$$

Because in the K_{sp} expression, OH^- is squared, the rule of thumb must reflect this operation.

Use 1×10^{-7} as $[\text{OH}^-]$, because $(1 \times 10^{-7} + 2x) \approx 1 \times 10^{-7}$ [x is so small compared to $1 \times 10^{-7} \text{ M}$].

$$K_{sp} = [\text{Cu}^{2+}][\text{OH}^-]^2$$

$$2.2 \times 10^{-20} = (x)(1 \times 10^{-7} + 2x)^2$$

$$2.2 \times 10^{-20} \approx (x)(1 \times 10^{-7})^2$$

$$x \approx 2.2 \times 10^{-6} \text{ M}$$

$$[\text{OH}^-] = 1 \times 10^{-7} \text{ M} + 2(2.2 \times 10^{-6} \text{ M}) \approx 1.0 \times 10^{-7} \text{ M}$$

$[\text{Cu}^{2+}] = 2.2 \times 10^{-6} \text{ M}$; $[\text{OH}^-] = 1 \times 10^{-7} \text{ M}$

- Selective Precipitation:** - by using a chemical reagent with an anion that will form precipitates with metal ions in a solution, the salts of each of these metal ions can be separated because one will precipitate first.
- when two or more precipitates is likely to form, **the solid with the lowest anion concentration at its K_{sp} will precipitate first.**
 - recall that **we can only use K_{sp} to compare solubilities if and only if the solids can produce the same number of ions.** Otherwise, we have to calculate the anion concentration needed of each salt individually.

Example 3: A chemical reagent consisting of $\text{SO}_4^{2-}(\text{aq})$ is slowly poured into a solution containing 3.00×10^{-3} mol/L of $\text{Ag}^+(\text{aq})$ and 1.50×10^{-4} mol/L of $\text{Pb}^{2+}(\text{aq})$. The K_{sp} values of Ag_2SO_4 and PbSO_4 are 1.2×10^{-5} and 1.3×10^{-8} respectively. Determine which of the above solids will precipitate first by calculating the $[\text{SO}_4^{2-}]$ require for each solid.



$[\text{SO}_4^{2-}] = 1.3 \text{ M}$ is needed for $\text{Ag}_2\text{SO}_4(\text{s})$ to form $[\text{SO}_4^{2-}] = 8.7 \times 10^{-5} \text{ M}$ is needed for $\text{PbSO}_4(\text{s})$ to form

Since $[\text{SO}_4^{2-}]$ is less for $\text{PbSO}_4(\text{s})$ to form, **lead (II) sulfate will precipitate first as $\text{SO}_4^{2-}(\text{aq})$ is slowly added to the solution. PbSO_4 can then be separated leaving the $\text{Ag}^+(\text{aq})$ in the filtrate.**

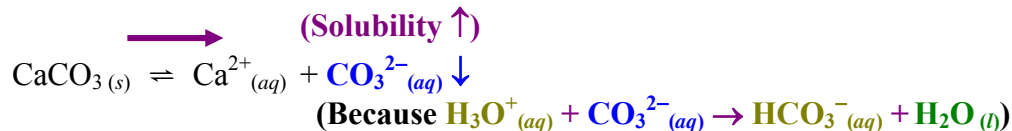
Assignment

16.8 pg. 742–744 #57 to 62, 97, 111

16.9: pH and Solubility

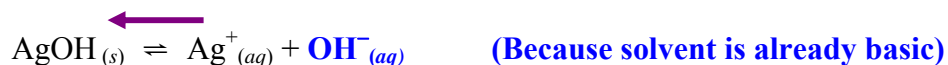
- pH and Solubility:** - when salts containing effective bases of weak acids (OH^- , S^{2-} , CO_3^{2-} , CH_3COO^- , and CrO_4^{2-}) are dissolved in acids, their solubilities increases.
- conversely, if these effective bases salts are dissolved in bases, their solubilities decreases.

Example: Calcium carbonate is dissolved in acidic solution.



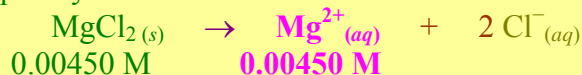
Example: Silver hydroxide is dissolved in basic solution.

(Solubility ↓)

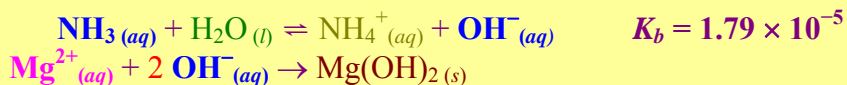


Example 1: Determine the minimum concentration of ammonia needed to start the precipitation of $\text{Mg}(\text{OH})_2_{(s)}$ from a 0.00450 M of $\text{MgCl}_2_{(aq)}$ solution. (K_b of $\text{NH}_3_{(aq)} = 1.79 \times 10^{-5}$ and K_{sp} of $\text{Mg}(\text{OH})_2_{(s)} = 1.2 \times 10^{-11}$)

MgCl_2 is very soluble and completely dissociates in water.



$\text{NH}_3_{(aq)}$ is a weak base and provides the $\text{OH}^-_{(aq)}$ (after the Brønsted-Lowry dissociation) for the formation of $\text{Mg}(\text{OH})_2_{(s)}$.



After $\text{Mg}(\text{OH})_2_{(s)}$ precipitated, some of it would dissociate back into ions according to K_{sp} .



Hence, we need to calculate the $[\text{OH}^-]$ first from the K_{sp} of $\text{Mg}(\text{OH})_2_{(s)}$. Then, we can use this $[\text{OH}^-]$ to find the original $[\text{NH}_3]$ required to initiate this precipitation.

$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2$$

$$1.2 \times 10^{-11} = (0.00450)[\text{OH}^-]^2$$

$$[\text{OH}^-] = \sqrt{\frac{1.2 \times 10^{-11}}{(0.00450)}} = 5.163977795 \times 10^{-5} \text{ M}$$

$[\text{OH}^-] = 5.16 \times 10^{-5} \text{ M}$ is needed for $\text{Mg}(\text{OH})_2_{(s)}$ to form.

Next, we can work backwards using the ICE box and K_b to find the $[\text{NH}_3]_0$.



	$[\text{NH}_3]$	$[\text{NH}_4^+]$	$[\text{OH}^-]$
Initial	$x \text{ M}$	0 M	0 M
Change	$- 5.163977795 \times 10^{-5} \text{ M}$	$+ 5.163977795 \times 10^{-5} \text{ M}$	$+ 5.163977795 \times 10^{-5} \text{ M}$
Equilibrium	$(x - 5.163977795 \times 10^{-5})$	$5.163977795 \times 10^{-5} \text{ M}$	$5.163977795 \times 10^{-5} \text{ M}$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \quad 1.79 \times 10^{-5} = \frac{(5.163977795 \times 10^{-5})^2}{(x - 5.163977795 \times 10^{-5})}$$

$$1.79 \times 10^{-5}(x - 5.163977795 \times 10^{-5}) = 2.666666667 \times 10^{-9}$$

$$1.79 \times 10^{-5}x - 9.24352025 \times 10^{-10} = 2.666666667 \times 10^{-9}$$

$$1.79 \times 10^{-5}x = 2.666666667 \times 10^{-9} + 9.24352025 \times 10^{-10}$$

$$x = \frac{3.591018692 \times 10^{-9}}{1.79 \times 10^{-5}} = 2.006155694 \times 10^{-4} \text{ M}$$

Minimum $[\text{NH}_3]_0 = x = 2.01 \times 10^{-4} \text{ mol/L}$

16.10: Complex Ion Equilibrium

Assignment

16.9 pg. 742 #63 to 68; pg. 745 #118

Complex Ion: - metal ion that is surrounded by Lewis Base (species with electron lone-pairs to donate).
- the resulting bonds between the metal ion and these Lewis Bases are called **ligands**.

Coordination Number: - the number of ligands that is attached to the metal ion.

Metal Ions and Coordination Numbers

Coordination Number	Metal Ions	Complex Ion Geometry
2	Cu ⁺ , Ag ⁺ , and Au ⁺	Linear
4	Al ³⁺ , Cu ⁺ , Au ⁺ , Mn ²⁺ , Co ²⁺ , Ni ²⁺ , Cu ²⁺ , Zn ²⁺ and Au ³⁺	Tetrahedral or Square Planar
6	Al ³⁺ , Mn ²⁺ , Fe ²⁺ , Co ²⁺ , Ni ²⁺ , Cu ²⁺ , Zn ²⁺ , Sc ³⁺ , Cr ³⁺ , Pt ³⁺ and Co ³⁺	Octahedral

Example 1: Write the net ionic equations of the following.

- a. A silver nitrate solution is mixed with an ammonia solution.

Complete Dissociation of Silver Nitrate: $\text{AgNO}_3(s) \rightarrow \text{Ag}^+(aq) + \text{NO}_3^-(aq)$

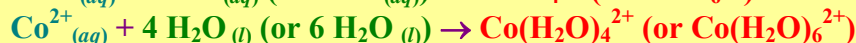
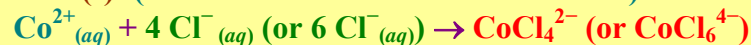
Net Ionic Equation: $\text{Ag}^+(aq) + 2 \text{NH}_3(aq) \rightarrow \text{Ag}(\text{NH}_3)_2^+$ (Ag⁺ has coordinate number 2)

- b. A cobalt (II) chloride solution is mixed with concentrated hydrochloric acid.

Complete Dissociation of Cobalt (II) Chloride: $\text{CoCl}_2(s) \rightarrow \text{Co}^{2+}(aq) + 2 \text{Cl}^-(aq)$

Hydrochloric Acid is a Strong Acid: $\text{HCl}(aq) \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq)$

Net Ionic Equation(s): (Co²⁺ has coordinate numbers 4 or 6)



Naming Complex Ions: - when naming a complex ion, first give the name(s) of the ligand(s), in alphabetical order, followed by the name of the metal.

Ligand (Lewis Base)	Name of Ligands	Metal Ions	Metal Name for Complex <u>Anion</u>
Water (H ₂ O)	Aqua	Iron (Fe)	Ferrate
Ammonia (NH ₃)	Ammine	Copper (Cu)	Cuprate
Methylamine (CH ₃ NH ₂)	Methylamine	Lead (Pb)	Plumbate
Carbon Monoxide (CO)	Carbonyl	Silver (Ag)	Argentate
Nitrogen Monoxide (NO)	Nitrosyl	Gold (Au)	Aurate
Hydroxide (OH ⁻)	Hydroxo	Tin (Sn)	Stannate
Cyanide (CN ⁻)	Cyano	Platinum (Pt)	Platinate
Thiocyanate (SCN ⁻)	Thiocyano	Cobalt (Co)	Cobaltate
Fluoride (F ⁻)	Fluoro	Aluminium (Al)	Aluminate
Chloride (Cl ⁻)	Chloro	Zinc (Zn)	Zincate
Bromide (Br ⁻)	Bromo	Nickel (Ni)	Nickelate
Iodide (I ⁻)	Iodo	Chromium (Cr)	Chrominate
Sulfate (SO ₄ ²⁻)	Sulfato	Scandium (Sc)	Scandiate
Sulfite (SO ₃ ²⁻)	Sulfito	Cadmium (Cd)	Cadmate
Nitrate (NO ₃ ⁻)	Nitrato		

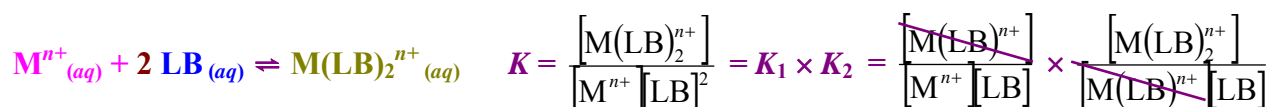
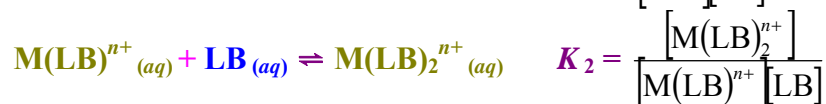
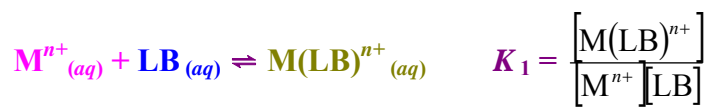
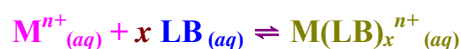
Nitrite (NO ₃ ⁻)	Nitrito		
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Special Notes on Naming Complex Ions:

- If a ligand is an anion whose name ends in *-ite* or *-ate*, the **final e** is changed to **o**.
(**Example:** change sulfat*e* to sulfat*o* and change nitrit*e* to nitrit*o*)
- If the ligand is an anion whose name ends in *-ide*, the **entire suffix ending** is changed to **o**.
(**Example.** change chlor*ide* to chlor*o* and cyan*ide* to cyan*o*)
- If the ligand is a neutral molecule, its common name is used. The important exceptions to this, however, are that water (H₂O) is called aqua, ammonia (NH₃) is called ammine, and carbon monoxide (CO) is called carbonyl.
- When there is **more than one of a particular ligand**, the number of ligands is designated by the appropriate **Greek prefix: di-, tri-, tetra-, penta-, hexa-, hepta-, etc.**
- If the **complex ion is an anion**, the **suffix -ate is added to the metal name**. The Latin name is often used for the metal in this case. For example, ferro rather than iron and cupro rather than copper.
- Following the name of the metal, the **oxidation number or original charge of the metal** is given using **Roman Numerals**.

Examples: Fe(CN)₆³⁻ - hexacyanoferrate (III) CoCl₄²⁻ - tetrachlorocobaltate (II)
 Al(H₂O)₆³⁺ - hexaaquaaluminium (III) Ni(SCN)₆⁴⁻ - hexathiocyanonickelate (II)
 Pt[(SO₄)₅Cl]³⁺ - pentasulfatochloroplatinum (III) Zn[(NO₂)₂(H₂O)₂] diaquadinitritozincate (II)

Formation Constants (K): - the equilibrium constant of the formation of complex ions from the metal ions and their Lewis bases.
 - also refer to as **Stability Constants**.
 - for a particular coordination number of a complex ion, there are equal number of formation constants.
 - the values of most formation constants are relatively large ($K_n \gg 1$). Therefore, we can **assume the formation of complex ions goes to completion**.

**Overall Formation Constant**

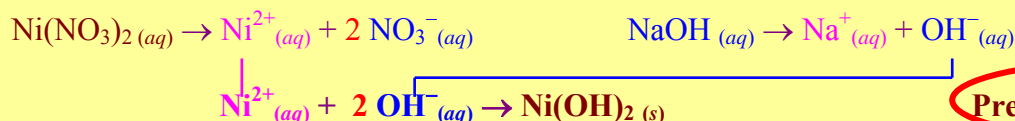
$$K = K_1 \times K_2 \times K_3 \times \dots \times K_x$$

Complex Ion and Solubility: - solubility of insoluble salts can be "re-dissolved" when mixed with sufficient Lewis base.

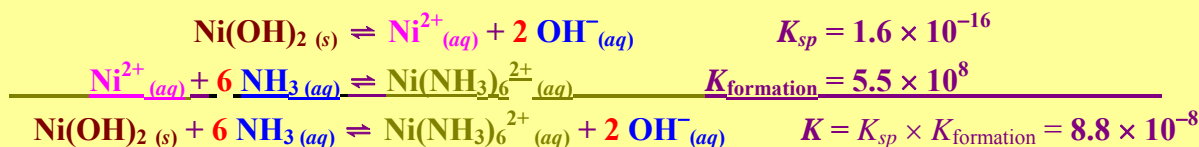
Example 3: A NaOH solution is added to $\text{Ni}(\text{NO}_3)_2(\text{aq})$ and a precipitate forms ($K_{sp} = 1.6 \times 10^{-16}$). Addition of 6.00 M of $\text{NH}_3(\text{aq})$ re-dissolved the precipitate into a complex ion with 6 as its coordination number.

- Identify the precipitate and the resulting complex ion.
- Calculate the solubility of the precipitate in $\text{NH}_3(\text{aq})$ if the overall formation constant is 5.5×10^8 . Neglect the volume of the NaOH and $\text{Ni}(\text{NO}_3)_2$ solutions.

a. Both $\text{Ni}(\text{NO}_3)_2$ and NaOH dissociates completely in water.



b. We need to calculate the overall formation constant for the complex ion from $\text{Ni}(\text{OH})_2(\text{s})$.



	$[\text{Ni}(\text{OH})_2]$	$[\text{NH}_3]$	$[\text{Ni}(\text{NH}_3)_6^{2+}]$	$[\text{OH}^-]$
Initial	----	6.00 M	0 M	$\approx 0 \text{ M}$
Change	----	$-6x$	$+x$	$+2x$
Equilibrium	----	$6 - 6x$	x	$\approx 2x$

We can assume $[\text{OH}]_0 \approx 0 \text{ M}$ because of the small value of K_{sp}

$$K = \frac{[\text{Ni}(\text{NH}_3)_6^{2+}][\text{OH}^-]^2}{[\text{NH}_3]^6}$$

$$8.8 \times 10^{-8} = \frac{(x)(2x)^2}{(6-6x)^6} = \frac{4x^3}{(6-6x)^6}$$

$$x = 0.085$$

```
solve(4x^3/(6-6x)^6-8.8e-8,x,0,(0,1))
.0845388877
```

Solubility of $\text{Ni}(\text{OH})_2$ in 6.00 M of $\text{NH}_3(\text{aq}) = 0.085 \text{ M}$

Assignment

16.10 pg. 742–743 #69 to 76, 89, 91, 98

16.10: Application of the Solubility Product to Qualitative Analysis

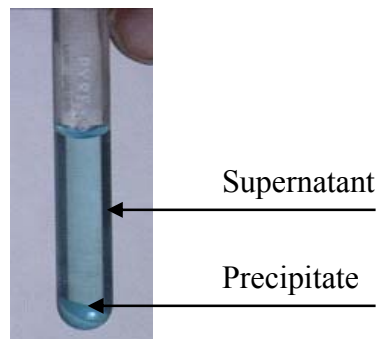
Centrifuge: - a device that speeds up the precipitation process by the use of centrifugal force (spinning the test tubes really fast).

Supernatant: - the liquid that remained after the precipitate is collected at the bottom of a test tube from the centrifuge process.

test tube wells

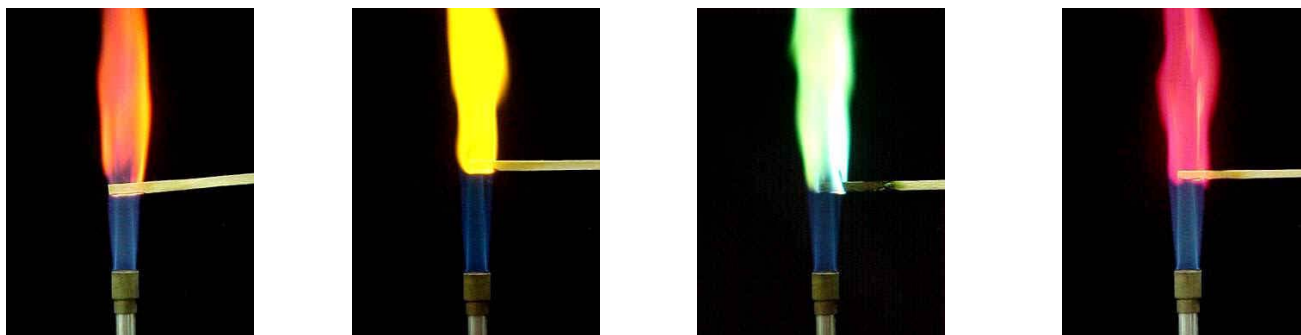


A typical laboratory centrifuge. Test tubes are placed inside the wells and the spinning will pull the precipitate down to the bottom.



Flame Test: - a test performed on a salt to detect the identity of a metal ion by the colour of the flame emitted as it is placed over a lighted Bunsen-burner.

- the principle of a metal ion emitting a certain visible colour is due to the electrons are excited to jump into higher orbitals as it is placed in a flame. As they come down to the lower orbital, a unique frequency is given in the visible spectrum producing a specific colour.



From left to right: LiCl , NaCl , CuCl_2 , and SrCl_2 produce different colors during the flame tests

Qualitative Analysis: - testing for the identities of metal ions by using selective precipitation or flame tests.
- usually both tests are used to identify the exact metal ions in the solution.

Developing a Qualitative Analysis Scheme for Cations:

1. **HCl is used to precipitate with Pb^{2+} , Hg^+ , Ag^+ , Tl^+ , and Cu^+** because their K_{sp} with Cl^- are small.
2. **Sulfide ion (S^{2-}) in acidic solution such as $\text{H}_2\text{S}_{(aq)}$ is commonly used to precipitate out the most insoluble salts such as (HgS , CuS , CdS , Bi_2S_3 , As_2S_3 , Sb_2S_3 , and SnS_2).** At low pH (high $[\text{H}^+]$), the concentration of S^{2-} is small. ($\text{H}_2\text{S}_{(aq)} \rightleftharpoons \text{H}^+_{(aq)} + \text{HS}^-_{(aq)}$ and $\text{HS}^-_{(aq)} \rightleftharpoons \text{H}^+_{(aq)} + \text{S}^{2-}_{(aq)}$) Therefore, metal-sulfide salts with higher K_{sp} values will remain in the solution as metal ions.

- Increasing the pH by adding NaOH will make the metal-sulfide salts with higher K_{sp} values to precipitate out (Mn^{2+} , Fe^{2+} , Ni^{2+} , Co^{2+} , Zn^{2+} , Fe^{3+} , Al^{3+} , and Cr^{3+}).** This is because a decrease in the $[H^+]$ will drive both acid dissociation reactions forward, increasing the $[S^{2-}]$.
- Finally, **CO_3^{2-} is added to precipitate the alkaline-earth ions such as Mg^{2+} , Ca^{2+} , Sr^{2+} and Ba^{2+} .**
- At each stage of the analysis, the residual (precipitate) can be filtered out and be subject to a flame test** for more accurate identification. Since alkali-metal ions are soluble with all anions, we will run a flame test on the filtrate or supernatant from step 4 to determine the identity of the remaining metal ions.

A Common Qualitative Analysis Scheme for most Metal Ions

Group Number	Solution Tested	Precipitating Reagent	Solids Precipitated	Metal Ions Detected	Flame Tests
1	Unknown	0.1 M HCl	PbCl ₂ , Hg ₂ Cl ₂ , AgCl, TlCl, and CuCl	Pb ²⁺ , Hg ⁺ , Ag ⁺ , Tl ⁺ , and Cu ⁺	Pb ²⁺ - light blue; Hg ⁺ - white; Ag ⁺ - gray; Tl ⁺ - green; Cu ⁺ - blue green
2	Filtrate or Supernatant from Group 1	H ₂ S at pH 1	HgS, CuS, CdS, Bi ₂ S ₃ , As ₂ S ₃ , Sb ₂ S ₃ , SnS ₂	Hg ²⁺ , Cu ²⁺ , Cd ²⁺ , Bi ³⁺ , As ³⁺ , Sb ³⁺ , and Sn ⁴⁺	Hg ²⁺ - white; Cu ²⁺ - blue green Cd ²⁺ - colourless; As ³⁺ - light blue; Bi ³⁺ - yellow brownish; Sb ³⁺ - green; Sn ⁴⁺ - colourless
3	Filtrate or Supernatant from Group 2	NaOH at pH 10	MnS, FeS, NiS, CoS, ZnS, Fe(OH) ₃ , Al(OH) ₃ , Cr(OH) ₃	Mn ²⁺ , Fe ²⁺ , Ni ²⁺ , Co ²⁺ , Zn ²⁺ , Fe ³⁺ , Al ³⁺ , and Cr ³⁺	Mn ²⁺ - violet Fe ²⁺ & Fe ³⁺ - yellow brownish red Ni ²⁺ - brown; Co ²⁺ - blue Zn ²⁺ - whitish green; Al ³⁺ - colourless; Cr ³⁺ - green
4	Filtrate or Supernatant from Group 3	Na ₂ CO ₃ at pH 10	MgCO ₃ , CaCO ₃ , SrCO ₃ , BaCO ₃	Mg ²⁺ , Ca ²⁺ , Sr ²⁺ and Ba ²⁺	Mg ²⁺ - white; Ca ²⁺ - yellowish red Sr ²⁺ - scarlet red; Ba ²⁺ - yellow green
5	Filtrate or Supernatant from Group 4	None	Soluble Ions	Li, Na ⁺ , K ⁺ , and NH ₄ ⁺	Li ⁺ - red; Na ⁺ - yellow; K ⁺ - violet; NH ₄ ⁺ - green

Assignment

16.10 pg. 743 #77, 79 to 82