

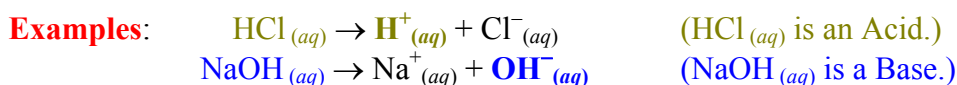
UNIT 8: ACIDS AND BASESChapter 20: Acids and Bases20.1: Describing Acids and BasesPhysical and Chemical Properties of Acid and Base

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).
Electrolytic (dissociate into ions in water)	Electrolytic (dissociate into ions in water)
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



**Acids:** - pure substance when dissolved in aqueous solution can dissociate a  $H^+$  ion (proton).

**Bases:** - pure substance when dissolved in aqueous solution can dissociate a  $OH^-$  ion (hydroxide ion).

Nomenclature of Acid

	Ionic Compound Name		Acid Name
1.	hydrogen ~ide	→	hydro~ic acid
2.	hydrogen ~ate	→	~ic acid
3.	hydrogen ~ite	→	~ous acid

**Example 1:** Name the following acids.

a.  $HI_{(aq)}$

hydrogen iodide

hydroiodic acid

b.  $HNO_3_{(aq)}$

hydrogen nitrate

nitric acid

c.  $HNO_2_{(aq)}$

hydrogen nitrite

nitrous acid

**Example 2:** Provide chemical formulas for the following acids.

a. hydrocyanic acid

*hydrocyanic acid*

hydrogen cyanide  $\Rightarrow \text{H}^+$  &  $\text{CN}^-$



b. oxalic acid

*oxalic acid*  $\Rightarrow$  hydrogen oxalate

$\text{H}^+$  &  $\text{OOC}\text{COO}^{2-}$



c. sulfurous acid

*sulfurous acid*

$\Rightarrow$  hydrogen sulfite

$\text{H}^+$  &  $\text{SO}_3^{2-}$



**Nomenclature of Base:** - most ionic bases come from metal cations (Group IA – Li, Na, K, Rb, Cs and some of Group IIA – Ca, Ba, Sr) with  $\text{OH}^-$  ion.  
- the naming of these ionic bases are “metal name” hydroxide.

**Example 3:** Name the following bases.

a.  $\text{KOH}_{(aq)}$

$\text{K}^+$  = potassium  
 $\text{OH}^-$  = hydroxide

potassium  
hydroxide

b.  $\text{Ba}(\text{OH})_{2(aq)}$

$\text{Ba}^{2+}$  = barium  
 $\text{OH}^-$  = hydroxide

barium  
hydroxide

**Example 4:** Provide chemical formulas for the following acids.

a. strontium hydroxide

$\text{Sr}^{2+}$  and  $\text{OH}^- \Rightarrow \text{Sr}(\text{OH})_2$

Need 1  $\text{Sr}^{2+}$  & 2  $(\text{OH}^-)$  to balance charges

b. rubidium hydroxide

$\text{Rb}^+$  and  $\text{OH}^- \Rightarrow \text{RbOH}$

Need 1  $\text{Rb}^+$  & 1  $\text{OH}^-$  to balance charges

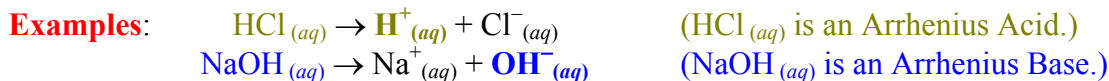
### Assignment

20.1: pg. 579 #1 to 5; pg. 609 #34

## 20.2: Hydrogen Ions and Acidity

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

**Arrhenius Concept:** - acids are  $\text{H}^+$  (proton) producers and bases are  $\text{OH}^-$  producers.

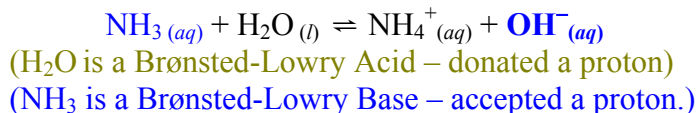
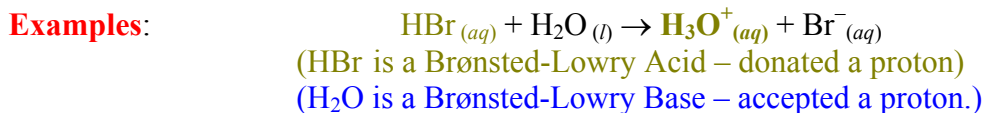


**Brønsted-Lowry Model:** - acids and bases **react with water** to **dissociate** where **acids are  $\text{H}^+$  (proton) donors** and **bases are  $\text{H}^+$  (proton) acceptors**.

- first proposed by Johannes Brønsted and Thomas Lowry.

**Hydronium Ion:** - an ion formed when an **acid “donated”  $\text{H}^+$  ion combined with a  $\text{H}_2\text{O}$  molecule** to form a  **$\text{H}_3\text{O}^+$  ion (hydronium ion)**.

- essentially has the same function as a  $\text{H}^+$  ion, but  $\text{H}_3\text{O}^+$  denotes that we are using the Brønsted-Lowry model.



**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)**,  **$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol/L}$**

### Autoionization of Water



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

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

$$\text{At pH} = 7: [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol/L}$$

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

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

$$2 \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{OH}^-_{(aq)}$$

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

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

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

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

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

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

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

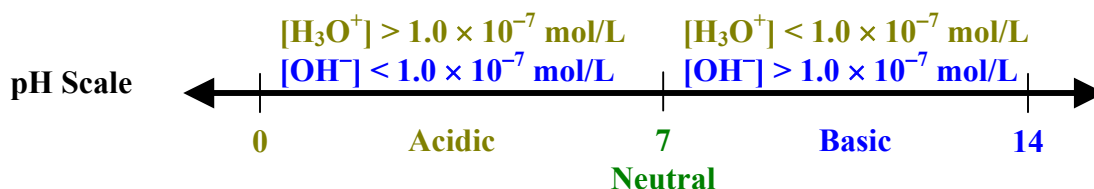
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

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

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

**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  $[\text{OH}^-]$** ).
- 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.



**Acidic Solution:** - solutions that has a **pH < 7**, where  **$[\text{H}_3\text{O}^+] > 1.0 \times 10^{-7} \text{ M}$**  and  **$[\text{OH}^-] < 1.0 \times 10^{-7} \text{ M}$** .

**Basic Solution:** - solutions that has a **pH > 7**, where  **$[\text{H}_3\text{O}^+] < 1.0 \times 10^{-7} \text{ M}$**  and  **$[\text{OH}^-] > 1.0 \times 10^{-7} \text{ M}$** .  
 - also known as **Alkaline Solution** (basic solutions that were from alkali metal hydroxides (LiOH, NaOH, KOH, RbOH, and CsOH) or alkaline metal hydroxides (Ca(OH)<sub>2</sub>, Ba(OH)<sub>2</sub>, and Sr(OH)<sub>2</sub>)

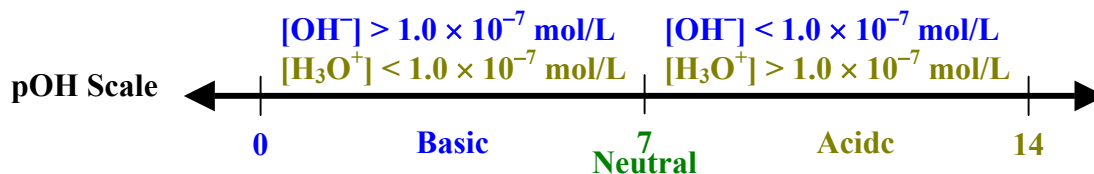
**Example:** pH of Some Common Substances

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 <sub>3</sub> )	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  $[\text{H}_3\text{O}^+]$ ).
  - 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 2 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.



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

**Example 3:** Calculate the pH and the pOH for the following solutions.

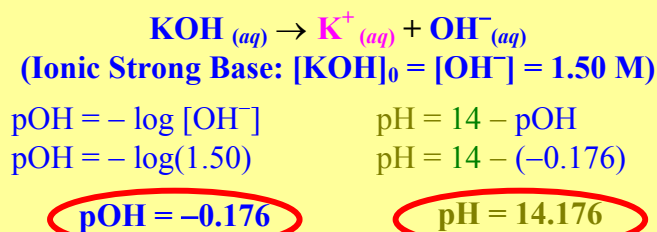
a.  $[\text{H}^+] = 0.0100 \text{ mol/L}$

$\text{pH} = -\log [\text{H}_3\text{O}^+]$	$\text{pOH} = 14 - \text{pH}$
$\text{pH} = -\log(0.0100)$	$\text{pOH} = 14 - 2.00$
<b>pH = 2.00</b>	<b>pOH = 12.00</b>

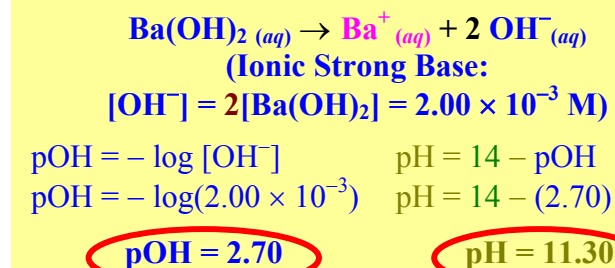
b.  $[\text{H}_3\text{O}^+] = 5.00 \times 10^{-6} \text{ M}$

$\text{pH} = -\log [\text{H}_3\text{O}^+]$	$\text{pOH} = 14 - \text{pH}$
$\text{pH} = -\log(5.00 \times 10^{-6})$	$\text{pOH} = 14 - 5.30$
<b>pH = 5.30</b>	<b>pOH = 8.70</b>

c. 1.50 M of  $\text{KOH}_{(aq)}$



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



**Example 4:** Calculate the  $[\text{H}_3\text{O}^+]$  and the  $[\text{OH}^-]$  for the following solutions.

a. pH = 4.00

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

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.00}$$

$$[\text{H}_3\text{O}^+] = 1.00 \times 10^{-4} \text{ M}$$

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

$$\text{pOH} = 14 - 4.00 = 10.00$$

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

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-10.00}$$

$$[\text{OH}^-] = 1.00 \times 10^{-10} \text{ M}$$

b. pOH = 3.00

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

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-3.00}$$

$$[\text{OH}^-] = 1.00 \times 10^{-3} \text{ M}$$

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

$$\text{pH} = 14 - 3 = 11.00$$

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

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-11.00}$$

$$[\text{H}_3\text{O}^+] = 1.00 \times 10^{-11} \text{ M}$$

c. pH = 12.83

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

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-12.83}$$

$$[\text{H}_3\text{O}^+] = 1.48 \times 10^{-13} \text{ M}$$

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

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

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

d. pOH = 9.67

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

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-9.67}$$

$$[\text{OH}^-] = 2.14 \times 10^{-10} \text{ M}$$

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

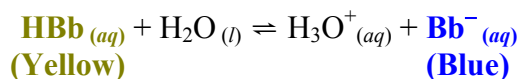
$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{2.14 \times 10^{-10}}$$

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

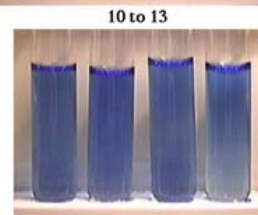
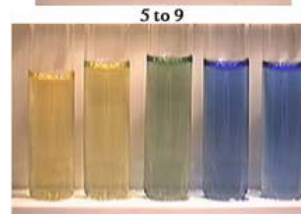
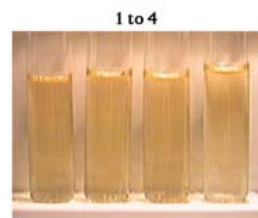
**Acid-Base Indicators:** - chemicals that change colours at a specific pH range.

- 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 has a different colour than its basic form.

**Example:** Bromothymol Blue ( $\text{HBb}_{(aq)} / \text{Bb}^-_{(aq)}$ ) changes colours from yellow to blue at a pH range of 6.0 to 7.6. At its **mid-range (6.8 – average of 6.0 and 7.6)**, the colour would be **green (yellow + blue)**.



**pH Meter:** - a device consists of an electrode that accurately measures pH.



Acid–Base Indicators at 298.15 K

Indicator	Suggested Abbreviation(s)	pH Range	Colour Change As pH Increases	$K_a$
methyl violet	$\text{HMv}_{(\text{aq})} / \text{Mv}^{-}_{(\text{aq})}$	0.0–1.6	yellow to blue	$\sim 10^{-1}$
cresol red	$\text{H}_2\text{Cr}_{(\text{aq})} / \text{HCr}^{-}_{(\text{aq})}$	0.0–1.0	red to yellow	$\sim 10^{-1}$
	$\text{HCr}_{(\text{aq})} / \text{Cr}^{2-}_{(\text{aq})}$	7.0–8.8	yellow to red	$3.5 \times 10^{-9}$
thymol blue	$\text{H}_2\text{Tb}_{(\text{aq})} / \text{HTb}^{-}_{(\text{aq})}$	1.2–2.8	red to yellow	$2.2 \times 10^{-2}$
	$\text{HTb}^{-}_{(\text{aq})} / \text{Tb}^{2-}_{(\text{aq})}$	8.0–9.6	yellow to blue	$6.3 \times 10^{-10}$
orange IV	$\text{HOr}_{(\text{aq})} / \text{Or}^{-}_{(\text{aq})}$	1.4–2.8	red to yellow	$\sim 10^{-2}$
methyl orange	$\text{HMo}_{(\text{aq})} / \text{Mo}^{-}_{(\text{aq})}$	3.2–4.4	red to yellow	$3.5 \times 10^{-4}$
bromocresol green	$\text{HBg}_{(\text{aq})} / \text{Bg}^{-}_{(\text{aq})}$	3.8–5.4	yellow to blue	$1.3 \times 10^{-5}$
methyl red	$\text{HMr}_{(\text{aq})} / \text{Mr}^{-}_{(\text{aq})}$	4.8–6.0	red to yellow	$1.0 \times 10^{-5}$
chlorophenol red	$\text{HCh}_{(\text{aq})} / \text{Ch}^{-}_{(\text{aq})}$	5.2–6.8	yellow to red	$5.6 \times 10^{-7}$
bromothymol blue	$\text{HBb}_{(\text{aq})} / \text{Bb}^{-}_{(\text{aq})}$	6.0–7.6	yellow to blue	$5.0 \times 10^{-8}$
phenol red	$\text{HPr}_{(\text{aq})} / \text{Pr}^{-}_{(\text{aq})}$	6.6–8.0	yellow to red	$1.0 \times 10^{-8}$
phenolphthalein	$\text{HPh}_{(\text{aq})} / \text{Ph}^{-}_{(\text{aq})}$	8.2–10.0	colourless to pink	$3.2 \times 10^{-10}$
thymolphthalein	$\text{HTh}_{(\text{aq})} / \text{Th}^{-}_{(\text{aq})}$	9.4–10.6	colourless to blue	$1.0 \times 10^{-10}$
alizarin yellow R	$\text{HAy}_{(\text{aq})} / \text{Ay}^{-}_{(\text{aq})}$	10.1–12.0	yellow to red	$6.9 \times 10^{-12}$
indigo carmine	$\text{Hlc}_{(\text{aq})} / \text{lc}^{-}_{(\text{aq})}$	11.4–13.0	blue to yellow	$\sim 10^{-12}$
1,3,5-trinitrobenzene	$\text{HNb}_{(\text{aq})} / \text{Nb}^{-}_{(\text{aq})}$	12.0–14.0	colourless to orange	$\sim 10^{-13}$

**Assignment**

**20.2:** pg. 582 #6 and 7; pg. 586 #8 and 9; pg. 587 #10 and 11; pg. 588 #12 and 13;  
pg. 589 #14 and 15; pg. 593 #16 to 18; pg. 609 #35 to 42

**20.3: Acid-Base Theories**

**Conjugate Base:** - the product formed after the Acid donated a  $H^+$ . (Acid  $\rightarrow$  Conjugate Base)  
 - behaves like a base when the reaction is looking from reverse.

**Conjugate Acid:** - the product formed after the Base accepted a  $H^+$ . (Base  $\rightarrow$  Conjugate Acid)  
 - behaves like an acid when the reaction is looking from reverse.

**Conjugate Acid-Base Pair:** - the (acid/conjugate base) or (base/conjugate acid) pairs.



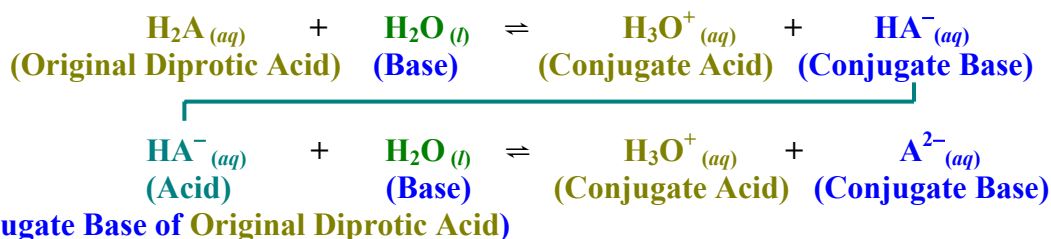
Conjugate Acid-Base Pairs:  $\text{HA}/\text{A}^-$  and  $\text{H}_2\text{O}/\text{H}_3\text{O}^+$



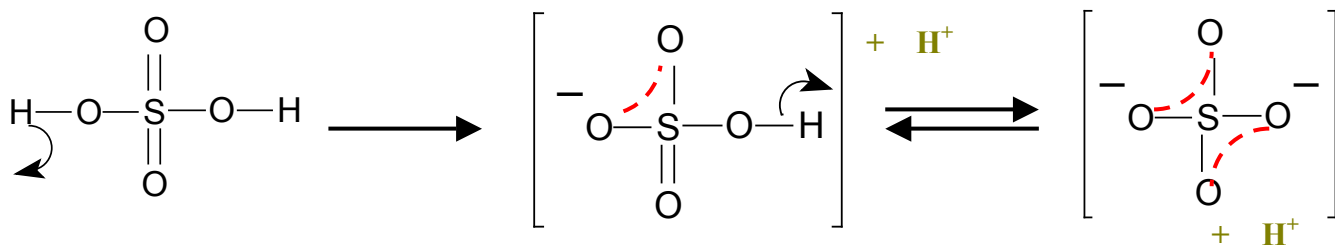
Conjugate Acid-Base Pairs:  $\text{B}/\text{HB}^+$  and  $\text{H}_2\text{O}/\text{OH}^-$

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

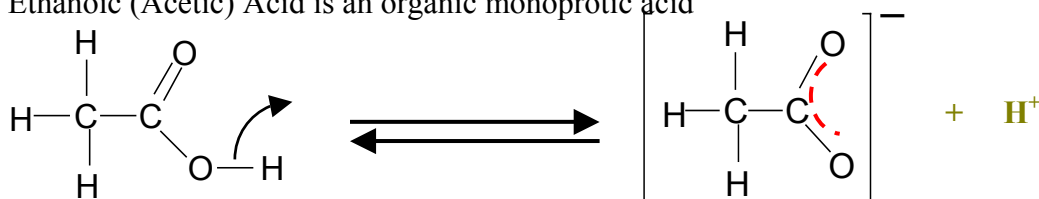
**Diprotic Acids:** - acids that can donate a maximum of two protons in stepwise dissociation.  
 - the conjugate base of the original acid can itself be an acid.  
 - just because a molecule has more than one proton does not mean it is diprotic (C-H bonds usually do not break, but some O-H bonds do).



**Example:** Sulfuric Acid ( $\text{H}_2\text{SO}_{4(aq)}$ ), successively donate protons to finally reach  $\text{SO}_4^{2-}_{(aq)}$



**Example:** Ethanoic (Acetic) Acid is an organic monoprotic acid





**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:  $\text{HOCCOO}^- (aq)$ ,  $\text{HSO}_4^- (aq)$ ,  $\text{HSO}_3^- (aq)$ ,  $\text{HCO}_3^- (aq)$ ,  $\text{HS}^- (aq)$ ,  $\text{HC}_6\text{H}_6\text{O}_6^- (aq)$ , and  $\text{H}_2\text{O} (l)$

**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:  $\text{H}_2\text{SO}_4 (aq)$ ,  $\text{HOCCOOH} (aq)$ ,  $\text{H}_2\text{SO}_3 (aq)$ ,  $\text{H}_2\text{C}_6\text{H}_6\text{O}_6 (aq)$ ,  $\text{H}_2\text{CO}_3 (aq)$ ,  $\text{H}_2\text{S} (aq)$

Diprotic Amphoteric Intermediates:  $\text{HSO}_4^- (aq)$ ,  $\text{HOCCOO}^- (aq)$ ,  $\text{HSO}_3^- (aq)$ ,  $\text{HC}_6\text{H}_6\text{O}_6^- (aq)$ ,  $\text{HCO}_3^- (aq)$ ,  $\text{HS}^- (aq)$

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:  $\text{H}_2\text{PO}_4^- (aq)$ ,  $\text{HPO}_4^{2-} (aq)$ ,  $\text{H}_2\text{AsO}_4^- (aq)$ ,  $\text{HAsO}_4^{2-} (aq)$ ,  $\text{H}_2\text{C}_6\text{H}_5\text{O}_7^- (aq)$ ,  $\text{HC}_6\text{H}_5\text{O}_7^{2-} (aq)$ ,  $\text{H}_2\text{BO}_3^- (aq)$ ,  $\text{HBO}_3^{2-} (aq)$

**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.

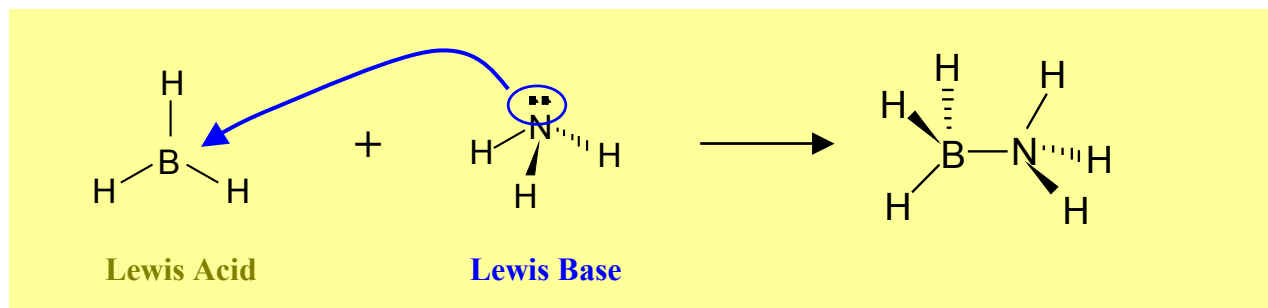
- especially suitable to explain metal hydrated cation as acid, or how metal oxides can turn into a base.

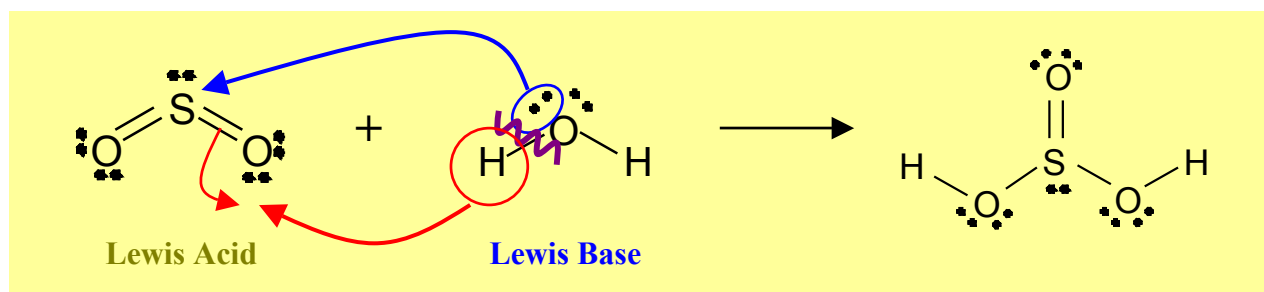
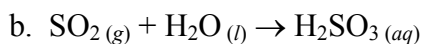
**Lewis Acid:** - a substance that ACCEPTS an Electron Pair (Lone-Pair).

**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.

a.  $\text{BH}_3 (g) + \text{NH}_3 (aq) \rightarrow \text{BH}_3\text{NH}_3 (aq)$





### Assignment

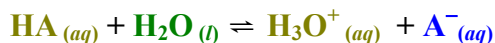
20.3: pg. 599 #19 to 25; pg. 609 #43 to 51

## 20.4A: Strengths of Acids and Bases

**Acid Dissociation Constant ( $K_a$ ):** - the equilibrium constant of a Brønsted-Lowry Acid Dissociation.

**Base Dissociation Constant ( $K_b$ ):** - the equilibrium constant of a Brønsted-Lowry Base Dissociation.

### Brønsted-Lowry Acid Dissociation



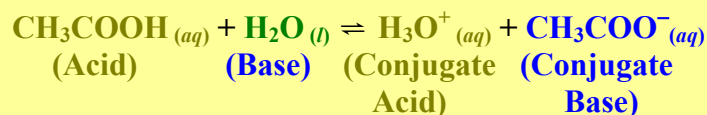
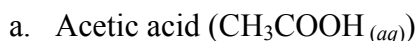
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

### Brønsted-Lowry Base Dissociation



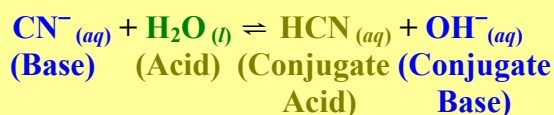
$$K_b = \frac{[\text{HB}^+][\text{OH}^-]}{[\text{B}]}$$

**Example 1:** Write the Brønsted-Lowry dissociation reaction of the following. Identify the Brønsted-Lowry acid and base, along with the conjugate acid and base. Determine the conjugate acid-base pairs. State the equilibrium expression of the dissociation reaction.



Conjugate Acid-Base Pairs:  
 $\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$  and  $\text{H}_2\text{O}/\text{H}_3\text{O}^+$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

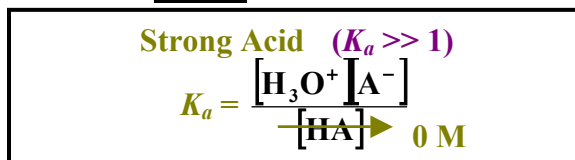


Conjugate Acid-Base Pairs:  
 $\text{CN}^-/\text{HCN}$  and  $\text{H}_2\text{O}/\text{OH}^-$

$$K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]}$$

**Strong Acids:** - acids that dissociate completely (100%) in water.

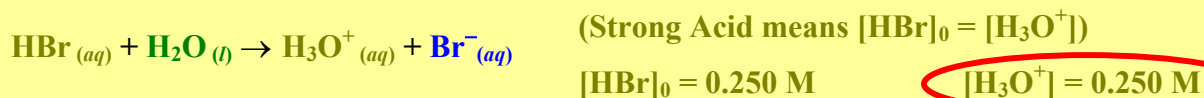
- when the acid dissociation constant ( $K_a$ ) is much greater than 1. ( $K_a \gg 1$ )
- the equilibrium position strongly favours the products.
- at equilibrium, the original acid concentration,  $[HA]_{eq} \approx 0$ ;  $[H_3O^+]_{eq} = [A^-]_{eq} = [HA]_0$ .
- the conjugate base,  $A^-$ , of a strong 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 <sub>3</sub> O <sup>+</sup> ]	[A <sup>-</sup> ]
Initial	x		0	0
Change	-x		+x	+x
Equilibrium	0		x	x

**Examples:** Strong Acids:  $HClO_4(aq)$ ,  $HI(aq)$ ,  $HBr(aq)$ ,  $HCl(aq)$ ,  $H_2SO_4(aq)$  and  $HNO_3(aq)$

**Example 2:** Write the dissociation reaction of 0.250 M of  $HBr(aq)$  and determine its  $[H_3O^+]$ .



**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 <sub>3</sub> O <sup>+</sup> ]	[A <sup>-</sup> ]
Initial	x		0	0
Change	-y (where $y \ll x$ )		+y	+y
Equilibrium	$(x - y) \approx x$		y	y

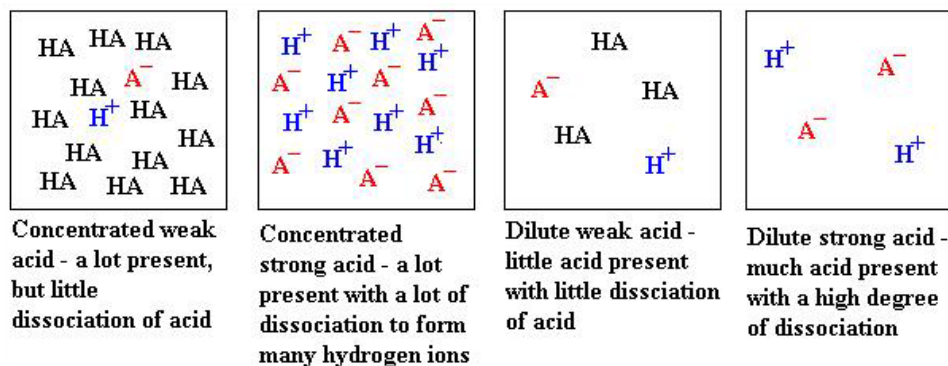
**Examples:** Some Weak Acids:  $HOOC-COOH(aq)$ ,  $H_2SO_3(aq)$ ,  $HSO_4^-(aq)$ ,  $H_3PO_4(aq)$ ,  $HNO_2(aq)$ ,  $H_3C_6H_5O_7(aq)$ ,  $HF(aq)$ ,  $HCOOH(aq)$ ,  $C_6H_8O_6(aq)$ ,  $C_6H_5COOH(aq)$ ,  $CH_3COOH(aq)$ ,  $H_2CO_3(aq)$ ,  $H_2S(aq)$ ,  $HOCl(aq)$ ,  $HCN(aq)$ ,  $NH_4^+(aq)$ , and  $H_3BO_3(aq)$

**$K_a$  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	$\text{HOOC}\text{COOH}(aq)$	$\text{HOOC}\text{COO}^-(aq)$	$5.6 \times 10^{-2}$
sulfurous acid ( $\text{SO}_2 + \text{H}_2\text{O}$ )	$\text{H}_2\text{SO}_3(aq)$	$\text{HSO}_3^-(aq)$	$1.4 \times 10^{-2}$
chlorous acid	$\text{HClO}_2(aq)$	$\text{ClO}_2^-(aq)$	$1.2 \times 10^{-2}$
hydrogen sulfate ion	$\text{HSO}_4^-(aq)$	$\text{SO}_4^{2-}(aq)$	$1.0 \times 10^{-2}$
phosphoric acid	$\text{H}_3\text{PO}_4(aq)$	$\text{H}_2\text{PO}_4^-(aq)$	$6.9 \times 10^{-3}$
nitrous acid	$\text{HNO}_2(aq)$	$\text{NO}_2^-(aq)$	$5.6 \times 10^{-3}$
arsenic acid	$\text{H}_3\text{AsO}_4(aq)$	$\text{H}_2\text{AsO}_4^-(aq)$	$5.0 \times 10^{-3}$
monochloroacetic acid	$\text{HC}_2\text{H}_2\text{ClO}_2$	$\text{C}_2\text{H}_2\text{ClO}_2^-(aq)$	$1.35 \times 10^{-3}$
citric acid	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7(aq)$	$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(aq)$	$7.4 \times 10^{-4}$
hydrofluoric acid	$\text{HF}(aq)$	$\text{F}^-(aq)$	$6.3 \times 10^{-4}$
methanoic (formic) acid	$\text{HCOOH}(aq)$	$\text{HCOO}^-(aq)$	$1.8 \times 10^{-4}$
hydrogen oxalate ion	$\text{HOOC}\text{COO}^-(aq)$	$\text{OOC}\text{COO}^{2-}(aq)$	$1.5 \times 10^{-4}$
lactic acid	$\text{HC}_3\text{H}_5\text{O}_3(aq)$	$\text{C}_3\text{H}_5\text{O}_3^-(aq)$	$1.38 \times 10^{-4}$
ascorbic acid (vitamin C)	$\text{H}_2\text{C}_6\text{H}_6\text{O}_6(aq)$	$\text{HC}_6\text{H}_6\text{O}_6^-(aq)$	$9.1 \times 10^{-5}$
benzoic acid	$\text{C}_6\text{H}_5\text{COOH}(aq)$	$\text{C}_6\text{H}_5\text{COO}^-(aq)$	$6.3 \times 10^{-5}$
ethanoic (acetic) acid	$\text{CH}_3\text{COOH}(aq)$	$\text{CH}_3\text{COO}^-(aq)$	$1.8 \times 10^{-5}$
dihydrogen citrate ion	$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^-(aq)$	$\text{HC}_6\text{H}_5\text{O}_7^{2-}(aq)$	$1.7 \times 10^{-5}$
hydrated aluminum (III) ion	$[\text{Al}(\text{H}_2\text{O})_6]^{3+}$	$[\text{AlOH}(\text{H}_2\text{O})_5]^{2+}$	$1.4 \times 10^{-5}$
propanoic acid	$\text{C}_2\text{H}_5\text{COOH}(aq)$	$\text{C}_2\text{H}_5\text{COO}^-(aq)$	$1.3 \times 10^{-5}$
carbonic acid ( $\text{CO}_2 + \text{H}_2\text{O}$ )	$\text{H}_2\text{CO}_3(aq)$	$\text{HCO}_3^-(aq)$	$4.5 \times 10^{-7}$
hydrogen citrate ion	$\text{HC}_6\text{H}_5\text{O}_7^{2-}(aq)$	$\text{C}_6\text{H}_5\text{O}_7^{3-}(aq)$	$4.0 \times 10^{-7}$
hydrosulfuric acid	$\text{H}_2\text{S}(aq)$	$\text{HS}^-(aq)$	$8.9 \times 10^{-8}$
dihydrogen arsenate ion	$\text{H}_2\text{AsO}_4^-(aq)$	$\text{HAsO}_4^{2-}(aq)$	$8.0 \times 10^{-8}$
hydrogen sulfite ion	$\text{HSO}_3^-(aq)$	$\text{SO}_3^{2-}(aq)$	$6.3 \times 10^{-8}$
dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^-(aq)$	$\text{HPO}_4^{2-}(aq)$	$6.2 \times 10^{-8}$
hypochlorous acid	$\text{HOCl}(aq)$	$\text{OCl}^-(aq)$	$4.0 \times 10^{-8}$
hypobromous acid	$\text{HOBr}(aq)$	$\text{OBr}^-(aq)$	$2.0 \times 10^{-9}$
hydrocyanic acid	$\text{HCN}(aq)$	$\text{CN}^-(aq)$	$6.2 \times 10^{-10}$
hydrogen arsenate ion	$\text{HAsO}_4^{2-}(aq)$	$\text{AsO}_4^{3-}(aq)$	$6.0 \times 10^{-10}$
boric acid	$\text{H}_3\text{BO}_3(aq)$	$\text{H}_2\text{BO}_3^-(aq)$	$5.8 \times 10^{-10}$
ammonium ion	$\text{NH}_4^+(aq)$	$\text{NH}_3(aq)$	$5.6 \times 10^{-10}$
phenol	$\text{C}_6\text{H}_5\text{OH}(aq)$	$\text{C}_6\text{H}_5\text{O}^-(aq)$	$1.6 \times 10^{-10}$
hydrogen carbonate ion	$\text{HCO}_3^-(aq)$	$\text{CO}_3^{2-}(aq)$	$4.7 \times 10^{-11}$
hypoiodous acid	$\text{HOI}(aq)$	$\text{OI}^-(aq)$	$2.0 \times 10^{-11}$
hydrogen ascorbate ion	$\text{HC}_6\text{H}_6\text{O}_6^-(aq)$	$\text{C}_6\text{H}_6\text{O}_6^{2-}(aq)$	$2.0 \times 10^{-12}$
hydrogen phosphate ion	$\text{HPO}_4^{2-}(aq)$	$\text{PO}_4^{3-}(aq)$	$4.8 \times 10^{-13}$
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

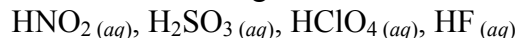


**Check out Strong and Weak Acids Animations at**  
<http://www.sgc.peachnet.edu/users/larnold/WWW/courses/1212/rev1212.html>

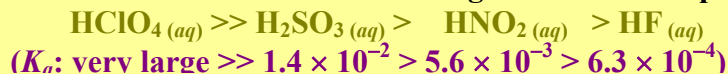
### Relative Strength of Acids and Conjugate Bases:

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 3:** Order the following acids from the strongest to the weakest.



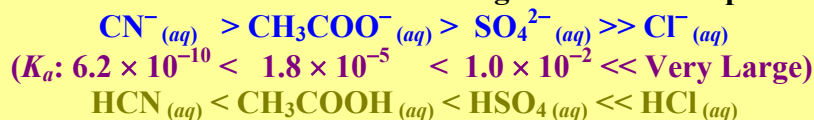
According to the Acid and Bases Relative Strength Table on the previous page:



**Example 4:** Order the following conjugate base from the strongest to the weakest.

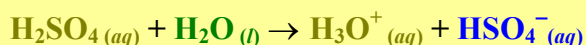


According to the Acids and Bases Relative Strength Table on the previous page:

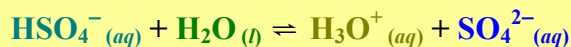


**Example 5:** Write the stepwise dissociation reaction for the following diprotic acids.

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

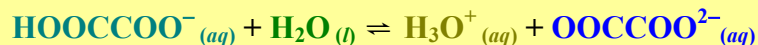
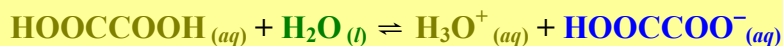


Direct Arrow because  $\text{H}_2\text{SO}_4(aq)$  is a Strong Acid ( $K_{a1}$  is very large)



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

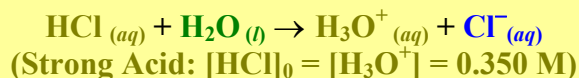
b.  $\text{HOCCOOH}(aq)$



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

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

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

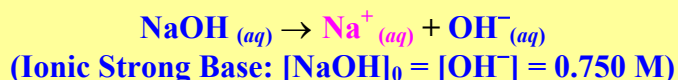


$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

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

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

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



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

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

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

Calculating the pH of Strong Acid or a Strong Base 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.

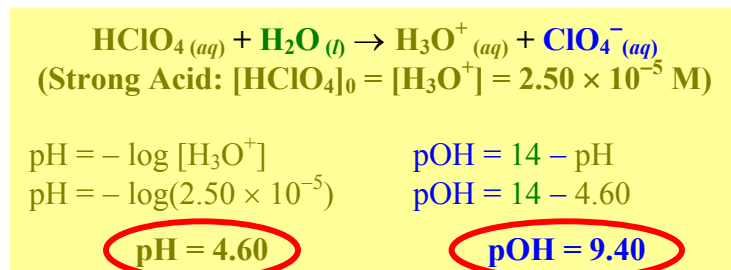
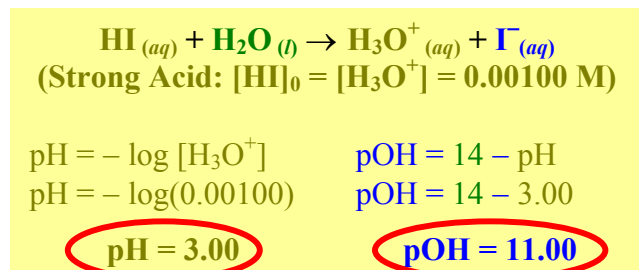
**Major Species of a Strong Acid:** - as strong acid dissociates completely ( $K_a > 1$ ) in water, the major species of all strong acids are  $\text{H}_3\text{O}^+$  and their conjugate bases.



**Example 7:** Calculate the pH and the pOH for the following solutions.

a. 0.00100 mol/L of  $\text{HI}_{(aq)}$

b.  $2.50 \times 10^{-5}$  M of  $\text{HClO}_{4(aq)}$



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

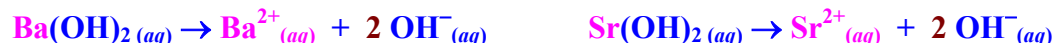
- all alkali bases (Group IA cations with  $\text{OH}^-$ ) and some alkaline bases (Group IIA cations with  $\text{OH}^-$ ) are considered as strong bases because they are ionic compound that dissociates completely.

**Examples:** Strong Ionic Bases:

- a. Alkali Bases:  $\text{LiOH}_{(aq)}$ ,  $\text{NaOH}_{(aq)}$ ,  $\text{KOH}_{(aq)}$ ,  $\text{RbOH}_{(aq)}$ , and  $\text{CsOH}_{(aq)}$   
(gives off 1 mole of  $\text{OH}^-$  when 1 mole of alkali base is dissolved)



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

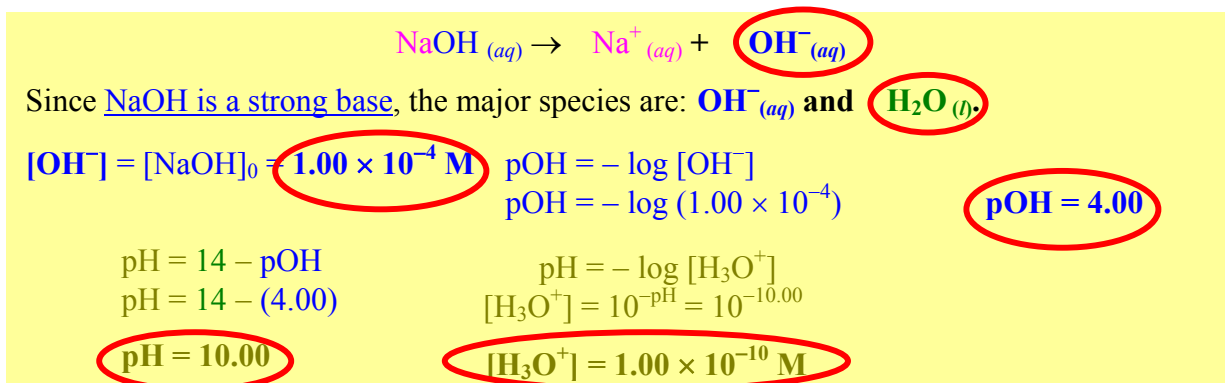


**Major Species of a Strong Base:** - as alkali- and alkaline bases dissociate completely in water, the major species of all strong bases are  $\text{OH}^-$  and water.

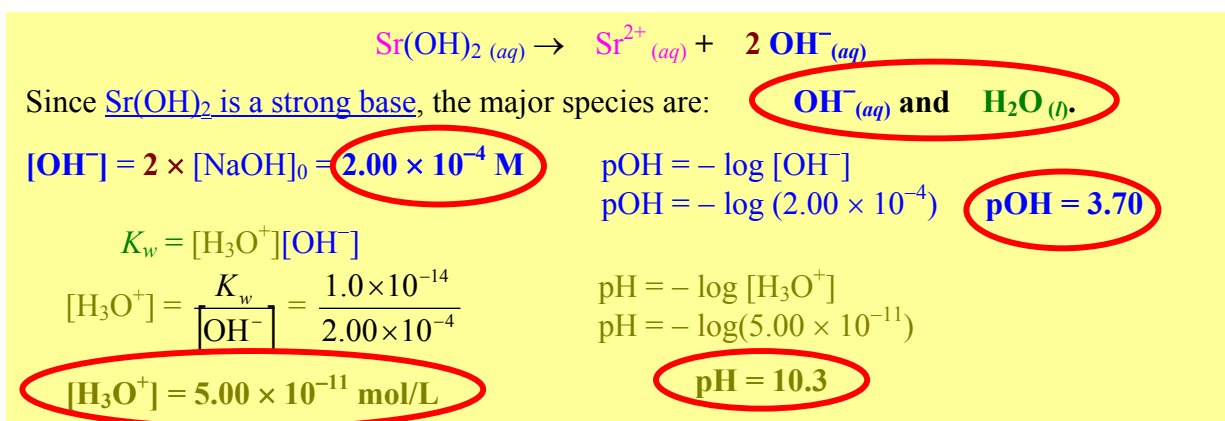
- special care must be taken with alkaline bases as they generate 2 moles of  $\text{OH}^-$  per 1 mole of solid dissolved.



**Example 8:** 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  $1.00 \times 10^{-4} \text{ M}$ .



**Example 9:** 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  $1.00 \times 10^{-4} \text{ M}$ .



### Calculating the $K_a$ of Weak Acid Solution

**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.

**Example 10:** 0.450 mol/L of an unknown weak monoprotic acid,  $\text{HA}_{(aq)}$  has a pH of 2.87. Determine the acid dissociation constant of this unknown acid.

**Major Species:**  $\text{HA}_{(aq)}$ ,  $\text{H}_2\text{O}_{(l)}$       **Strongest Acid:**  $\text{HA}$  ( $K_a = ?$ )

$\text{HA}$  is a weak acid and undergoes Brønsted-Lowry Dissociation.

$$\text{HA}_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{A}^-_{(aq)}$$

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

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-2.87}$$

$$[\text{H}_3\text{O}^+] = 0.0013489629 \text{ M} = [\text{A}^-]$$

	$[\text{HA}]$	$[\text{H}_3\text{O}^+]$	$[\text{A}^-]$
<b>Initial</b>	0.450 M	0	0
<b>Change</b>	-0.0013489629 M	+0.0013489629 M	+0.0013489629 M
<b>Equilibrium</b>	0.4486510371 M	0.0013489629 M	0.0013489629 M

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{(0.0013489629)(0.0013489629)}{(0.4486510371)}$$

$$K_a = 4.06 \times 10^{-6}$$

**Assignment: 20.4A: pg. 605 #26 to 33; pg. 609 #52 to 56**



**20.4B: Acidity and Basicity of Weak Acids and Bases**

**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  $H_3O^+$ .

**Using the SOLVE function of the TI-83 Plus Calculator:**

1. Manipulate the Equation so one-side is equal to 0.

**Example:**  $2.4 \times 10^{-2} = \frac{x^2}{(0.4-x)}$  becomes  $0 = \frac{x^2}{(0.4-x)} - 2.4 \times 10^{-2}$

2. Callout the **solve** function.

**solve (Equation Expression, X, guess, {lower boundary, upper boundary})**

Diagram illustrating the TI-83 Plus calculator steps:

- 2<sup>nd</sup> CATALOG** → **0** → **,** → **X,T,θ,n** → **2<sup>nd</sup> {** → **Enter 0** → **Enter 0** → **Enter Initial Concentration of Weak Acid  $[HA]_0$**  → **2<sup>nd</sup> }** → **)**
- ALPHA** → **e<sup>x</sup>** → **S** → **LN** → **ENTER**
- Scroll Down from S** → **ENTER** (Select solve)
- ENTER Parameters** → **ENTER**

Calculator screens showing the process:

- Screen 1:** CATALOG, sin(, sin-1(, sinh(, sinh-1(, SinReg, solve(, SortA(
- Screen 2:** solve(X<sup>2</sup>/(0.4-X) - 2.4E-2, X, 0, 0, 0, .4)
- Screen 3:** solve(X<sup>2</sup>/(0.4-X) - 2.4E-2, X, 0, 0, 0, .4) = .0867117014  
Value of x

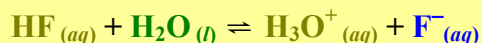


**Example 1:** Determine the  $[\text{H}_3\text{O}^+]$ , and pH of 0.200 mol/L of  $\text{HF}_{(aq)}$  ( $K_a = 6.3 \times 10^{-4}$ ).

**Major Species:**  $\text{HF}_{(aq)}$ ,  $\text{H}_2\text{O}_{(l)}$

**Strongest Acid:**  $\text{HF}$  ( $K_a = 6.3 \times 10^{-4}$ ) ( $K_a$  for  $\text{H}_2\text{O}$  is  $1.0 \times 10^{-14}$ )

$\text{HF}$  is a weak acid and undergoes Brønsted-Lowry Dissociation.



	$[\text{HF}]$	$[\text{H}_3\text{O}^+]$	$[\text{F}^-]$
<b>Initial</b>	0.200 M	0	0
<b>Change</b>	$-x$	$+x$	$+x$
<b>Equilibrium</b>	$(0.2 - x)$	$x$	$x$

**CANNOT use Approximation:**

$$\frac{[\text{HF}]_0}{K_a} = \frac{0.200 \text{ M}}{6.3 \times 10^{-4}} = 317.5 < 1000$$

Have to use  $(0.2 - x)$  in the denominator

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} \quad 6.3 \times 10^{-4} = \frac{(x)(x)}{(0.2 - x)} = \frac{x^2}{(0.2 - x)}$$

$$6.3 \times 10^{-4} (0.2 - x) = x^2$$

$$(1.26 \times 10^{-4}) - (6.3 \times 10^{-4})x = x^2$$

$$x^2 + (6.3 \times 10^{-4})x - (1.26 \times 10^{-4}) = 0$$

(Quadratic Equation: Apply the Quadratic Formula!)

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \quad a = 1 \quad b = 6.3 \times 10^{-4} \quad c = -1.26 \times 10^{-4}$$

$$x = \frac{-(6.3 \times 10^{-4}) \pm \sqrt{(6.3 \times 10^{-4})^2 - 4(1)(-1.26 \times 10^{-4})}}{2(1)}$$

$$x = 0.0109143911$$

$$x = -0.0115443911 \text{ (omit negative } x)$$

$$[\text{H}_3\text{O}^+] = 0.0109 \text{ mol/L}$$

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

$$\text{pH} = -\log(0.0109)$$

$$\text{pH} = 1.96$$

```
solve(X^2/(.2-X)-
6.3E-4,X,0,(0,0.
2))
.0109143911
```

OR

```
solve(X^2+6.3E-4X
-1.26E-4,X,0,(0,
0.2))
.0109143911
```

**Percent Dissociation:** - the amount of  $[\text{H}_3\text{O}^+]$  dissociated from the original  $[\text{HA}]_0$  expressed in percentage.

- strong acids will have % dissociation = 100%.
- weak acids will have % Dissociation < 100%

#### Percent Dissociation of Acids

$$\% \text{ Dissociation} = \frac{[\text{H}_3\text{O}^+]}{[\text{HA}]_0} \times 100\%$$

**Example 2:** 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<sub>(aq)</sub>, H<sub>2</sub>O<sub>(l)</sub>

**Strongest Acid:** HA ( $K_a = ?$ ) ( $K_a$  for H<sub>2</sub>O is  $1.0 \times 10^{-14}$  and has a 0.000 01 % dissociation)

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



	[HA]	[H <sub>3</sub> O <sup>+</sup> ]	[A <sup>-</sup> ]
Initial	0.0500 M	0	0
Change	$-1.9 \times 10^{-4}$	$+1.9 \times 10^{-4}$	$1.9 \times 10^{-4}$
Equilibrium	0.04981 M	$1.9 \times 10^{-4}$ M	$1.9 \times 10^{-4}$ M

$$\% \text{ Dissociation} = \frac{[\text{H}_3\text{O}^+]}{[\text{HA}]_0} \times 100\%$$

$$[\text{H}_3\text{O}^+] = (\% \text{Dissociation})[\text{HA}]_0 / 100\%$$

$$[\text{H}_3\text{O}^+] = (0.38\%)(0.0500 \text{ M}) / 100\%$$

$$[\text{H}_3\text{O}^+] = 1.9 \times 10^{-4} \text{ M} = [\text{A}^-]$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{(1.9 \times 10^{-4})(1.9 \times 10^{-4})}{(0.04981)}$$

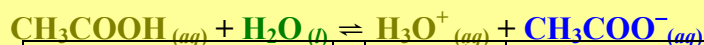
$$K_a = 7.2 \times 10^{-7}$$

**Example 3:** Determine the [H<sub>3</sub>O<sup>+</sup>] and pH of 0.200 mol/L of CH<sub>3</sub>COOH<sub>(aq)</sub> ( $K_a = 1.8 \times 10^{-5}$ ). Calculate the percent dissociation of this acid.

**Major Species:** CH<sub>3</sub>COOH<sub>(aq)</sub>, H<sub>2</sub>O<sub>(l)</sub>

**Strongest Acid:** CH<sub>3</sub>COOH ( $K_a = 1.8 \times 10^{-5}$ )

CH<sub>3</sub>COOH is a weak acid and undergoes Brønsted-Lowry Dissociation.



	[CH <sub>3</sub> COOH]	[H <sub>3</sub> O <sup>+</sup> ]	[CH <sub>3</sub> COO <sup>-</sup> ]
Initial	0.200 M	0	0
Change	-x	+x	+x
Equilibrium	(0.2 - x)	x	x

**CAN use Approximation:**

$$\frac{[\text{CH}_3\text{COOH}]_0}{K_a} = \frac{0.200 \text{ M}}{1.8 \times 10^{-5}} = 11111 \geq 1000$$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \quad 1.8 \times 10^{-5} = \frac{(x)(x)}{(0.2 - x)} \approx \frac{x^2}{(0.2)}$$

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

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

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

$$x \approx 0.00190$$

$$[\text{H}_3\text{O}^+] = 0.00190 \text{ mol/L}$$

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

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

$$\text{pH} = 2.72$$

$$\% \text{ Dissociation} = \frac{[\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]_0} \times 100\%$$

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

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

**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,  $\text{HB}^+$ , of a weak base is itself a stronger weak acid (can easily donate protons to do the reverse reaction).

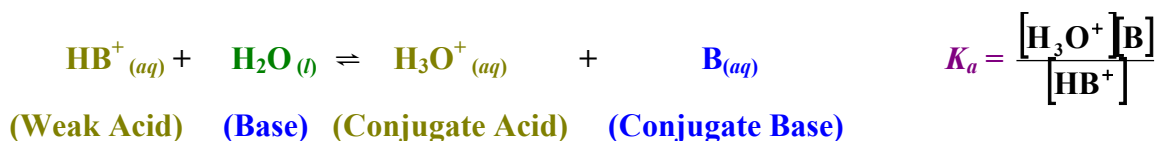
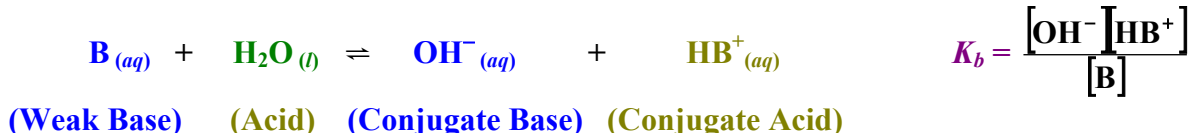


	[B]		[OH <sup>-</sup> ]	[HB <sup>+</sup> ]
Initial	x		0	0
Change	-y (where y << x)		+y	+y
Equilibrium	(x - y) ≈ x		y	y

**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**

$$K_w = K_a \times K_b$$

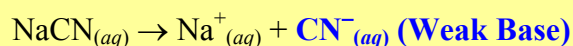
**Major Species of a Weak Base:** - since weak acid do not dissociate completely ( $K_b < 1$ ) in water, the major species of all weak bases are their original form and water.

Procedure to calculate pH of Weak Base Solutions:

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

**Example 4:** Determine the  $[OH^-]$ ,  $[H_3O^+]$ , pOH, pH and % dissociation of 0.200 mol/L of  $NaCN_{(aq)}$ . ( $K_a$  of  $HCN = 6.2 \times 10^{-10}$ ).

**NaCN dissociates completely in water:**

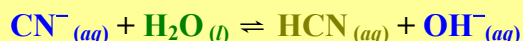


**Major Species:**  $CN^-_{(aq)}$ ,  $H_2O_{(l)}$

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{6.2 \times 10^{-10}} \quad K_b = 1.613 \times 10^{-5} \quad \text{(take a few more decimal places to avoid round off errors)}$$

**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.



**CAN use Approximation:**

$$\frac{[CN^-]_0}{K_b} = \frac{0.200 \text{ M}}{1.613 \times 10^{-5}} = 12399 \geq 1000$$

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

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

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

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

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

$$x \approx 0.0018$$

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

$$pOH = -\log [OH^-]$$

$$pOH = -\log(0.0018)$$

$$pOH = 2.75$$

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

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

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

$$pH = 14 - pOH$$

$$pH = 14 - (2.75)$$

$$pH = 11.25$$

$$pH = -\log [H_3O^+]$$

$$[H_3O^+] = 10^{-pH} = 10^{-11.25}$$

$$[H_3O^+] = 5.6 \times 10^{-12} \text{ M}$$

### Assignment

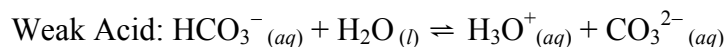
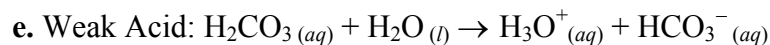
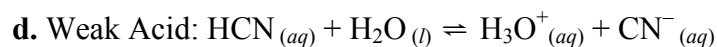
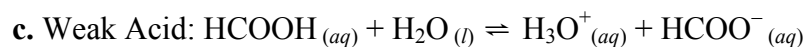
### Acids and Bases Worksheet

Acids and Bases Worksheet

1. What is the pH of  $1.00 \times 10^{-4}$  M of  $\text{HClO}_{4(aq)}$ ?
2. What is the pH of 0.00300 M of  $\text{HNO}_{3(aq)}$ ?
3. A potassium hydroxide solution is prepared by dissolving 6.00 g of KOH in 1.50 L of water. What are the pOH and the pH of the solution?
4. A solution was made by dissolving 0.637 g of  $\text{Ba}(\text{OH})_2$  in 200 mL final volume. What is the pOH and pH of the solution?
5. A solution of HBr has a pH of 2.50. What was the initial concentration of  $\text{HBr}_{(aq)}$ ?
6. A solution of  $\text{Sr}(\text{OH})_2$  has a measured pH of 11.70. What is the initial molar concentration of  $\text{Sr}(\text{OH})_2$  in the solution?
7. Identify the following as strong or weak acid. Write the dissociation equation for each of these acids. For any polyprotic acids, write all dissociation steps.  
a.  $\text{HNO}_2$       b.  $\text{H}_2\text{SO}_4$       c.  $\text{HCOOH}$       d.  $\text{HCN}$       e.  $\text{H}_2\text{CO}_3$
8. For each of the weak acids in question 7, write the appropriate  $K_a$  expression.
9. Identify the following as strong or weak base. Write the dissociation equation for each of these bases after any ionic salt dissolve completely into their cations and anions.  
a. NaF      b. KOCl      c.  $\text{Ba}(\text{OH})_2$       d.  $\text{NaC}_6\text{H}_5\text{COO}$       e.  $\text{NH}_3$
10. For each of the weak bases in question 9, write the appropriate  $K_b$  expression.
11. Calculate the  $K_b$  values of the weak bases in question 9.
12. Calculate the pH and the percent dissociation of the following solutions.  
a. 0.200 M of benzoic acid      b. 0.0500 M of methanoic acid
13. Lactic acid ( $K_a = 1.38 \times 10^{-4}$ ) is a weak monoprotic acid that builds up in muscle during exercises causing fatigue. Determine the pH and the percent dissociation of 0.500 M of lactic acid,  $\text{HC}_3\text{H}_5\text{O}_3(aq)$ .
14. Calculate the pH and the percent dissociation of the following solutions.  
a. 0.300 M of KOCl      b. 0.0800 M of  $\text{NH}_3$
15. Quinine, an important drug in treating malaria, is a weak Brønsted-Lowry base that we can represent as  $Qu$ . Determine the pH and the percent dissociation of  $6.50 \times 10^{-3}$  mol/L of quinine.  $\text{HQu}^+$ , the conjugate acid of  $Qu$ , has a  $K_a$  value of  $5.62 \times 10^{-9}$ .

Answers

1. pH = 4
2. pH = 2.52
3. pOH = 1.15; pH = 12.85
4. pOH = 1.73; pH = 12.27
5.  $[\text{HBr}]_0 = 0.00316$  M
6.  $[\text{Sr}(\text{OH})_2]_0 = 0.00251$  M
- 7a. Weak Acid:  $\text{HNO}_{2(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{NO}_2(aq)$   
b. Strong Acid:  $\text{H}_2\text{SO}_{4(aq)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{H}_3\text{O}^+_{(aq)} + \text{HSO}_4^-(aq)$   
Weak Acid:  $\text{HSO}_4^-(aq) + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{SO}_4^{2-}(aq)$



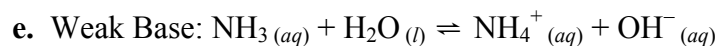
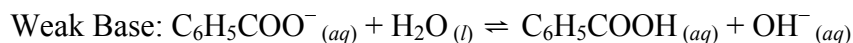
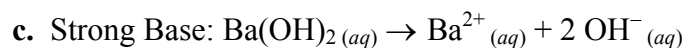
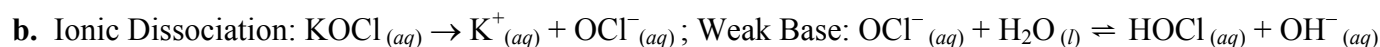
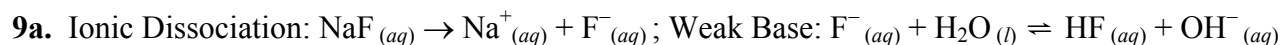
8a.  $K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$

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

c.  $K_a = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$

d.  $K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]}$

e.  $K_a = \frac{[\text{H}_3\text{O}^+][\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$ ;  $K_a = \frac{[\text{H}_3\text{O}^+][\text{CO}_3^{2-}]}{[\text{HCO}_3^-]}$



10a.  $K_b = \frac{[\text{HF}][\text{OH}^-]}{[\text{F}^-]}$

b.  $K_b = \frac{[\text{HOCl}][\text{OH}^-]}{[\text{OCl}^-]}$

c. Strong Base (No  $K_b$  expression)

d.  $K_b = \frac{[\text{C}_6\text{H}_5\text{COOH}][\text{OH}^-]}{[\text{C}_6\text{H}_5\text{COO}^-]}$

e.  $K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$

11a.  $K_b = 1.6 \times 10^{-11}$

b.  $K_b = 2.5 \times 10^{-7}$

c. Strong Base ( $K_b$  is very large)

d.  $K_b = 1.6 \times 10^{-10}$

e.  $K_b = 1.8 \times 10^{-5}$

12a. pH = 2.45; % Dissociation = 1.77%

b. pH = 2.54; % Dissociation = 5.82%

13. pH = 2.08; % Dissociation = 1.66%

14a. pH = 10.44; % Dissociation = 0.0913%

b. pH = 11.08; % Dissociation = 1.49%

15. pH = 10.03; % Dissociation = 1.65%

## Chapter 21: Neutralization

### 21.1: Neutralization

**Neutralization:** - the reaction of an acid and a base to form water and an acid-base salt.

**Acid-Base Salt:** - simply known as **salt** – mostly soluble ionic compound left after from a **neutralization** reaction.

- for strong acid and strong base neutralization, it consists of the anions of a strong acid and a cations of a strong base.

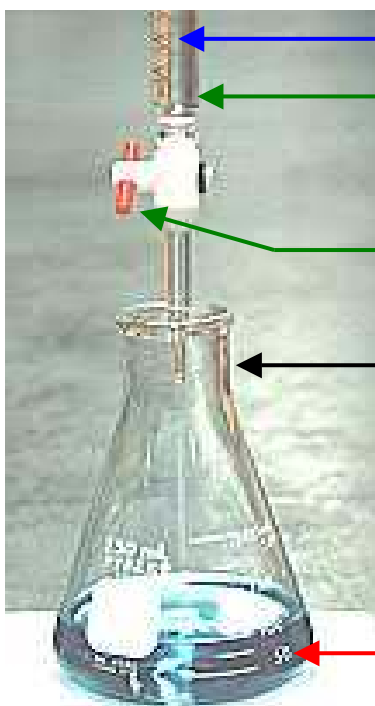
#### Different Forms of Strong Acid and Strong Base Neutralization Chemical Equation.

1. Molecular Equation:  $\text{HA}_{(aq)} + \text{MOH}_{(aq)} \rightarrow \text{HOH}_{(l)} + \text{MA}_{(aq)}$  ← Acid-Base Salt
2. Complete Ionic Equation:  $\text{H}^+_{(aq)} + \text{A}^-_{(aq)} + \text{M}^+_{(aq)} + \text{OH}^-_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)} + \text{M}^+_{(aq)} + \text{A}^-_{(aq)}$
3. Net-Ionic Equation:  $\text{H}^+_{(aq)} + \text{OH}^-_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)}$

**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  $OH^-$ . ( $n_{H^+} = n_{OH^-}$ )

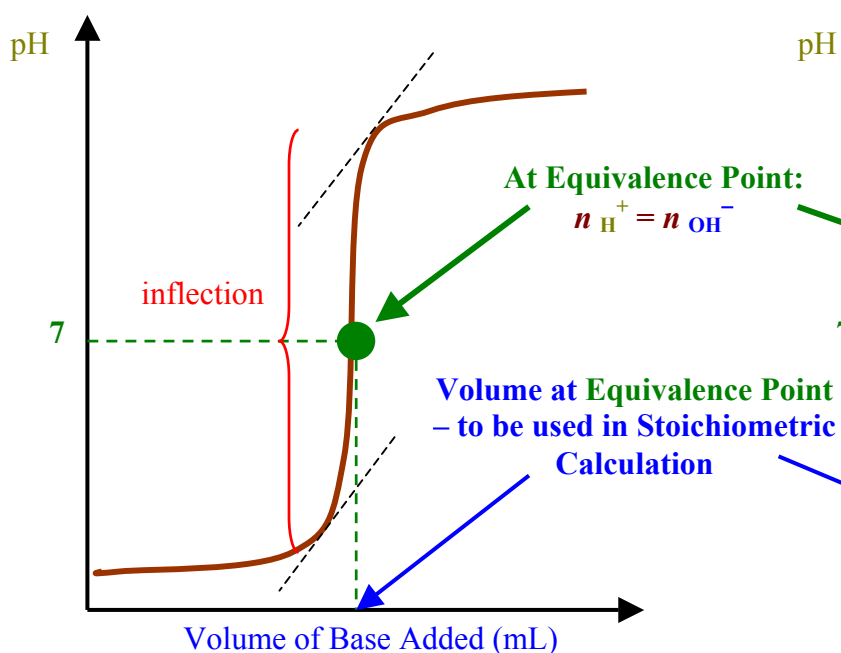
**Endpoint:** - a point where the indicator actually changes colour to indicate neutralization is completed.

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

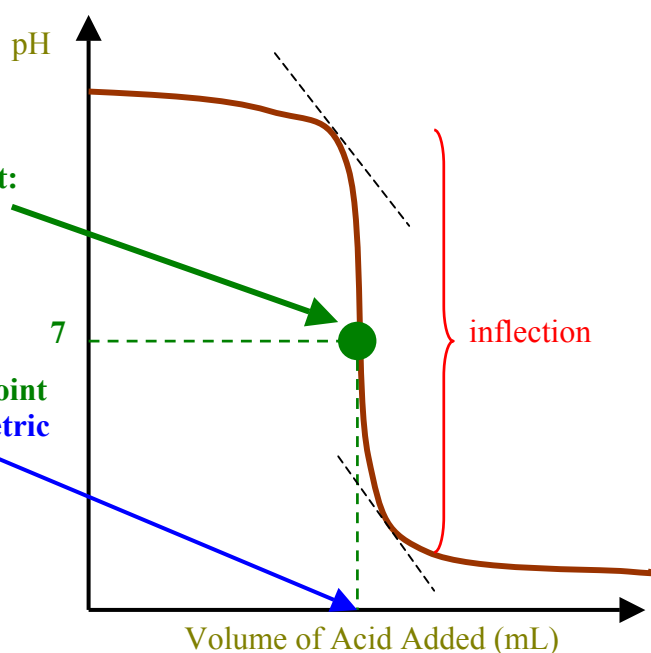
**Inflection:** - the part of the curve where there is a sudden rise or drop in pH.  
- the midpoint of the inflection marks the equivalence point.

**Titration Between Strong Acids and Strong Base:** - **Equivalence Point** always occur at **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  $HCl_{(aq)}$  is titrated with 0.200 M of  $KOH_{(aq)}$  at:

- 0 mL of  $KOH_{(aq)}$  added.
- 5.00 mL of  $KOH_{(aq)}$  added.
- 15.00 mL of  $KOH_{(aq)}$  added.
- 20.00 mL of  $KOH_{(aq)}$  added.

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

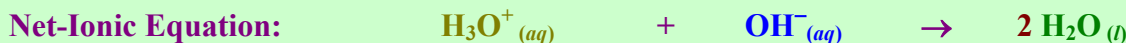


$$pH = -\log [H_3O^+] = -\log(0.100)$$

$$pH = 1.00$$



b. At 5.00 mL of KOH<sub>(aq)</sub> added: (Before Stoichiometric Point)



	$n$ of $\text{H}_3\text{O}^+$	$n$ of $\text{OH}^-$
Before	$(0.100 \text{ mol/L})(30.00 \text{ mL}) = 3.00 \text{ mmol}$	$(0.200 \text{ mol/L})(5.00 \text{ mL}) = 1.00 \text{ mmol (LR)}$
Change	$-1.00 \text{ mmol}$	$-1.00 \text{ mmol}$
After	$2.00 \text{ mmol}$	$0.00 \text{ mmol}$

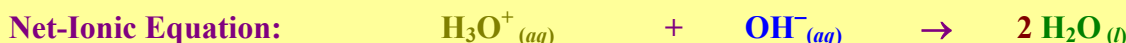
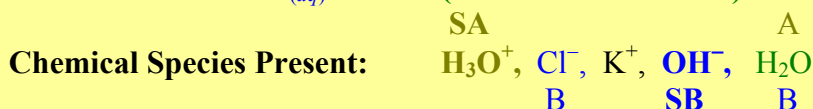
$$[\text{H}_3\text{O}^+] = \frac{n_{\text{H}_3\text{O}^+}}{\text{Total Volume}} = \frac{2.00 \text{ mmol}}{(30.00 \text{ mL} + 5.00 \text{ mL})} = 0.05714 \text{ mol/L}$$

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

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

$$\text{pH} = 1.24$$

c. At 15.00 mL of KOH<sub>(aq)</sub> added: (Stoichiometric Point)



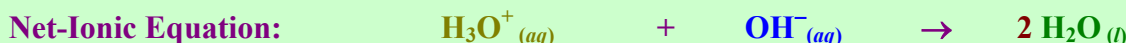
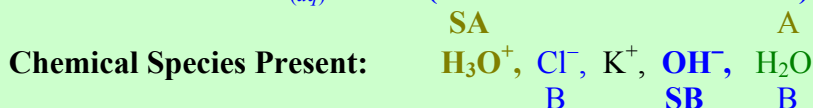
	$n$ of $\text{H}_3\text{O}^+$	$n$ of $\text{OH}^-$
Before	$(0.100 \text{ mol/L})(30.00 \text{ mL}) = 3.00 \text{ mmol}$	$(0.200 \text{ mol/L})(15.00 \text{ mL}) = 3.00 \text{ mmol}$
Change	$-3.00 \text{ mmol}$	$-3.00 \text{ mmol}$
After	$0.00 \text{ mmol}$	$0.00 \text{ mmol}$

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

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

$$\text{pH} = 7.00$$

d. At 20.00 mL of KOH<sub>(aq)</sub> added: (After Stoichiometric Point)



	$n$ of $\text{H}_3\text{O}^+$	$n$ of $\text{OH}^-$
Before	$(0.100 \text{ mol/L})(30.00 \text{ mL}) = 3.00 \text{ mmol}$	$(0.200 \text{ mol/L})(20.00 \text{ mL}) = 4.00 \text{ mmol}$
Change	$-3.00 \text{ mmol}$ (Limiting Reagent)	$-3.00 \text{ mmol}$
After	$0.00 \text{ mmol}$	$1.00 \text{ mmol}$

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

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

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

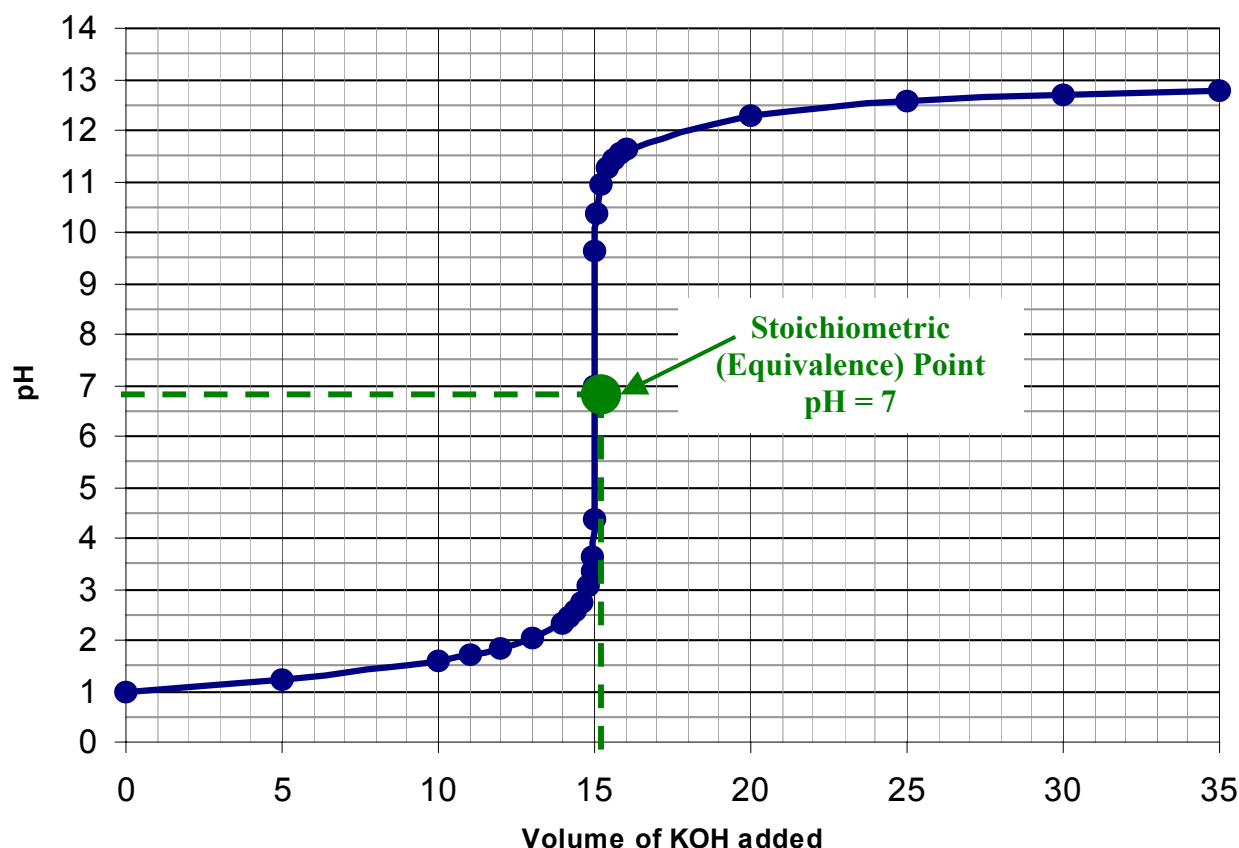
(pH's results from similar Calculations with other volumes after the stoichiometric point are listed below.)

$$\text{pOH} = 1.70$$

$$\text{pH} = 12.30$$

**pH of 30.0 mL of 0.100 M HCl<sub>(aq)</sub> titrated by 0.200 M of KOH<sub>(aq)</sub>**

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	<b>15.00 mL (Equivalence)</b>	<b>7.00</b>	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
14.40 mL	2.57	15.20 mL	10.95		

**pH of 30.0 mL of 0.100 M HCl titrated by 0.200 M of KOH**

**Standardized Solution:** - **a solution prepared** by dissolving an accurately measured amount of solute or precisely diluted **stock solution** (the original chemical reagent from the chemical manufacturers) **shortly before the experiment**.

- a standardized solution is needed because the concentration of some chemical reagents changes over time.

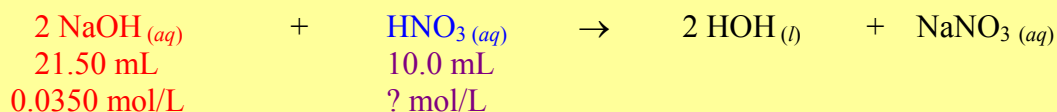
**Example:**  $\text{H}_2\text{CO}_3(\text{aq})$  tends to dissociate into  $\text{H}_2\text{O}(\text{l})$  and  $\text{CO}_2(\text{g})$  over time. Therefore, a given concentration of  $\text{H}_2\text{CO}_3(\text{aq})$  from 2 weeks ago will not have the stated concentration for the experiment today.

**Example 1:** Use the following observation table to determine the concentration of sulfuric acid.

10.0 mL of $\text{HNO}_3(aq)$ titrated by 0.0350 mol/L of $\text{NaOH}(aq)$				
	Trial 1	Trial 2	Trial 3	Trial 4
Initial Volume	0.32 mL	24.19 mL	3.48 mL	24.97 mL
Final Volume	24.19 mL	45.71 mL	24.97 mL	46.47 mL
Volume of NaOH added	23.87 mL	21.52 mL	21.49 mL	21.50 mL
Bromothymol Blue Colour	Blue	Green	Green	Green

First, we have to complete the table by subtracting the final and the initial volumes. Since the titration is completed when the indicator turns green, we only average the result of the last 3 trials.

$$\text{Average Volume of KOH added} = \frac{21.52 \text{ mL} + 21.49 \text{ mL} + 21.50 \text{ mL}}{3} = 21.50 \text{ mL}$$



$$\textcircled{1} n_{\text{NaOH}} = CV = (0.0350 \text{ mol/L}) (21.50 \text{ mL}) = 0.7525 \text{ mmol}$$

$$\textcircled{2} n_{\text{HNO}_3} = 0.7525 \text{ mmol NaOH} \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol NaOH}} = 0.7525 \text{ mmol HNO}_3$$

$$\textcircled{3} C_{\text{HNO}_3} = \frac{n}{V} = \frac{0.7525 \text{ mmol}}{10.0 \text{ mL}} = 0.07525 \text{ mol/L}$$

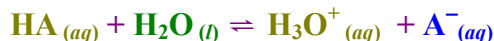
$$[\text{HNO}_3] = 0.0753 \text{ mol/L} = 75.3 \text{ mmol/L}$$

### Assignment

21.1 pg. 616 #1 and 2; pg. 618 #3 and 4; pg. 624 #17 to 19;  
pg. 640 #36 to 39; pg. 641 # 62

**21.2: Salts in Solutions**

**Buffered Solution:** - a solution that resists a change in pH when a small amount of  $\text{H}_3\text{O}^+$  or  $\text{OH}^-$  is added.  
 - consists of a pair of weak acid/conjugate base common ion or a pair of weak base/conjugate acid common ion.

**Acidic Buffered Solution:**

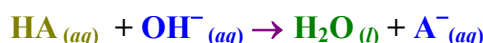
a. Small Amounts of  $\text{H}_3\text{O}^+$  is Added:



( $\text{H}^+$  - Strong Acid reacts completely with  $\text{A}^-$ )

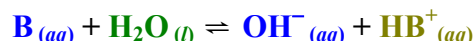
(More HA – Weak Acid: pH will only be lowered SLIGHTLY!)

b. Small Amounts of  $\text{OH}^-$  is Added:



( $\text{OH}^-$  - Strong Base reacts completely with HA)

(More  $\text{A}^-$  – Weak Base: pH will only be raised SLIGHTLY!)

**Basic Buffered Solution:**

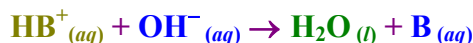
a. Small Amounts of  $\text{H}_3\text{O}^+$  is Added:



( $\text{H}^+$  - Strong Acid reacts completely with B)

(More  $\text{HB}^+$  – Weak Acid: pH will only be lowered SLIGHTLY!)

b. Small Amounts of  $\text{OH}^-$  is Added:



( $\text{OH}^-$  - Strong Base reacts completely with  $\text{HB}^+$ )

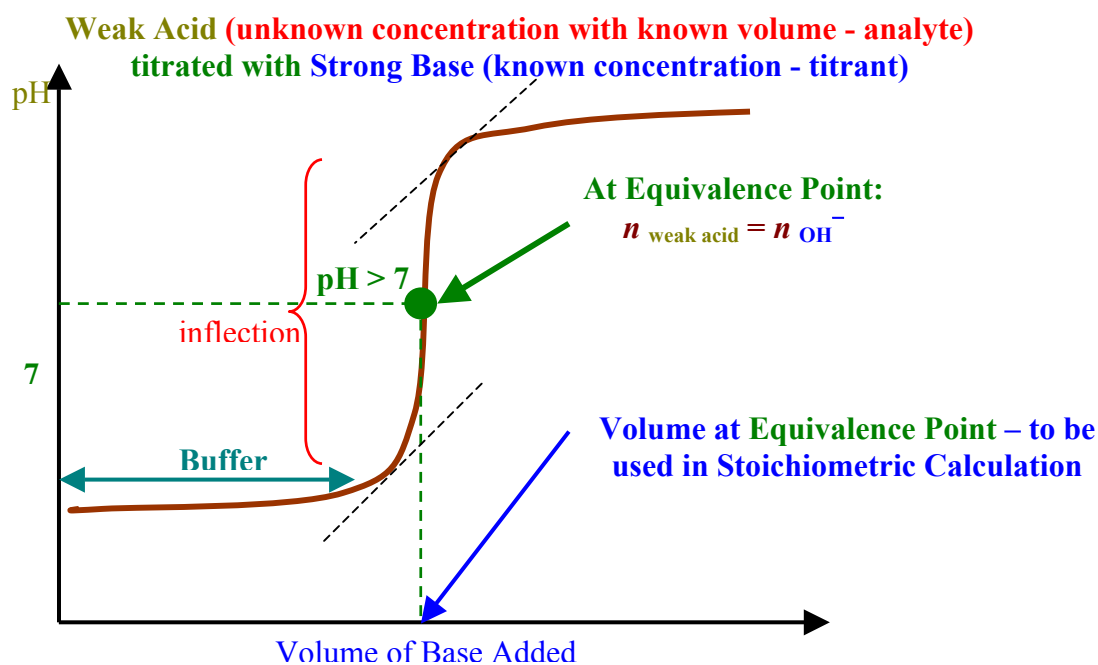
(More B – Weak Base: pH will only be raised SLIGHTLY!)

**Example:**  $\text{HCO}_3^-_{(aq)} / \text{CO}_3^{2-}$  is a buffer because  $\text{HCO}_3^-$  acts as an acid and  $\text{CO}_3^{2-}$  acts as a base whenever a small amount of strong base or acid is added to it. Thus, pH will remain somewhat constant.

When small amounts of strong base is added:  $\text{HCO}_3^-_{(aq)} + \text{OH}^-_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)} + \text{CO}_3^{2-}_{(aq)}$

When small amounts of strong acid is added:  $\text{CO}_3^{2-}_{(aq)} + \text{H}_3\text{O}^+_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)} + \text{HCO}_3^-_{(aq)}$

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



At 15.00 mL of  $\text{KOH}_{(aq)}$  added: **(Stoichiometric Point)**

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

	<i>n</i> of HCOOH	<i>n</i> of OH <sup>-</sup>	<i>n</i> of HCOO <sup>-</sup>
<b>Before</b>	(0.100 mol/L)(30.00 mL) = 3.00 mmol	(0.200 mol/L)(15.00 mL) = 3.00 mmol	0 mmol
<b>Change</b>	- 3.00 mmol	- 3.00 mmol	+ 3.00 mmol
<b>After</b>	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  $\text{OH}^-$  and  $\text{HCOOH}$  are used up, the resulting solution follows the regular weak base dissociation.

	$\text{HCOO}^-_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{HCOOH}_{(aq)} + \text{OH}^-_{(aq)}$		
	$[\text{HCOO}^-]$		$[\text{HCOOH}]$ $[\text{OH}^-]$
Initial	0.0667 M	0	0 M
Change	$-x$	$+x$	$+x$
Equilibrium	$(0.0667 - x)$	$x$	$x$

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

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

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

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

$$\frac{[\text{HCOOH}]_0}{K_b} = \frac{0.0667 \text{ M}}{5.556 \times 10^{-11}} = 1.2 \times 10^9 \geq 1$$

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

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

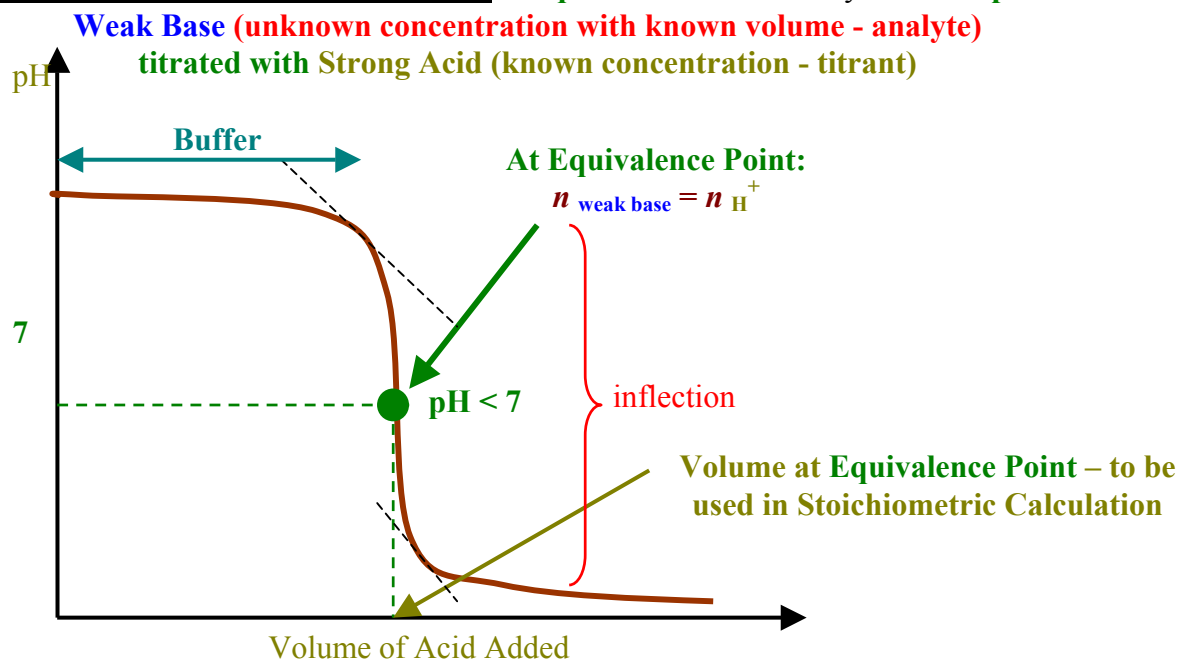
**pH = 8.28**

**pH of 30.0 mL of 0.100 M  $\text{HCOOH}_{(aq)}$  titrated by 0.200 M of  $\text{KOH}_{(aq)}$**

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	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	<b>15.00 mL (Equivalence)</b>	<b>8.28</b>	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



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



**Example 3:** Calculate the pH when 30.0 mL of 0.100 M of methylamine,  $\text{CH}_3\text{NH}_2(aq)$  ( $K_b = 4.38 \times 10^{-4}$ ) is titrated with 15.0 mL of 0.200 mol/L  $\text{HCl}(aq)$  at its equivalence (stoichiometric) point.

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

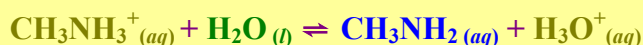
Chemical Species Present:  $\text{CH}_3\text{NH}_2$ ,  $\text{H}_3\text{O}^+$ ,  $\text{Cl}^-$ ,  $\text{H}_2\text{O}$   
SB SA A/B

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 $\text{CH}_3\text{NH}_2$	$n$ of $\text{H}_3\text{O}^+$	$n$ of $\text{CH}_3\text{NH}_3^+$
Before	$(0.100 \text{ mol/L})(30.00 \text{ mL})$ $= 3.00 \text{ mmol}$	$(0.200 \text{ mol/L})(15.00 \text{ mL})$ $= 3.00 \text{ mmol}$	0 mmol
Change	$- 3.00 \text{ mmol}$	$- 3.00 \text{ mmol}$	$+ 3.00 \text{ mmol}$
After	0 mmol	0 mmol	3.00 mmol

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

At Stoichiometric Point, after all the  $\text{H}_3\text{O}^+$  and  $\text{CH}_3\text{NH}_2$  are used up, the resulting solution follows the regular weak acid dissociation.



	$[\text{CH}_3\text{NH}_3^+]$	$[\text{CH}_3\text{NH}_2]$	$[\text{H}_3\text{O}^+]$
Initial	0.0667 M	0	0 M
Change	$-x$	$+x$	$+x$
Equilibrium	$(0.0667 - x)$	$x$	$x$

$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{4.38 \times 10^{-4}} \quad K_a = 2.28 \times 10^{-11} = \frac{[\text{CH}_3\text{NH}_2][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{NH}_3^+]}$$

CAN use Approximation:

$$\frac{[\text{CH}_3\text{NH}_3^+]_0}{K_a} = \frac{0.0667 \text{ M}}{2.28 \times 10^{-11}} = 2.9 \times 10^9 \geq 1000$$

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

$$2.28 \times 10^{-11} = \frac{(x)(x)}{(0.0667 - x)} \approx \frac{x^2}{0.0667} \quad [\text{H}_3\text{O}^+] = x = 1.23 \times 10^{-6} \text{ mol/L}$$

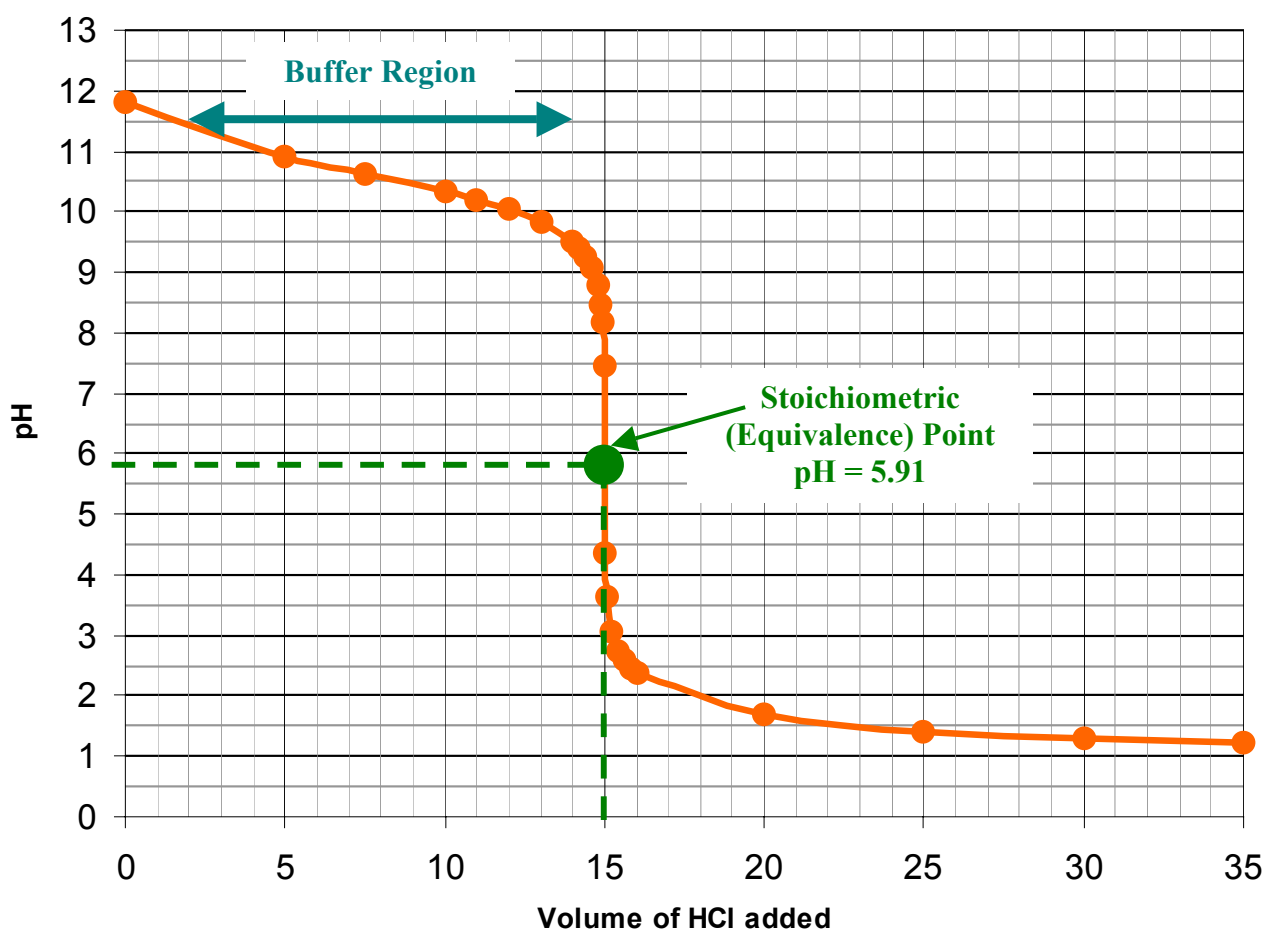
$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(1.23 \times 10^{-6})$$

$$\text{pH} = 5.91$$

**pH of 30.0 mL of 0.100 M  $\text{CH}_3\text{NH}_2(aq)$  titrated by 0.200 M of  $\text{HCl}(aq)$**

Volume of HCl added	pH	Volume of HCl added	pH	Volume of HCl added	pH
0.00 mL	11.81	14.40 mL	9.26	15.20 mL	3.05
5.00 mL	10.92	14.60 mL	9.08	15.40 mL	2.75
7.50 mL	10.63	14.80 mL	8.77	15.60 mL	2.58
10.00 mL	10.33	14.90 mL	8.47	15.80 mL	2.46
11.00 mL	10.20	14.95 mL	8.16	16.00 mL	2.36
12.00 mL	10.03	14.99 mL	7.46	20.00 mL	1.70
13.00 mL	9.82	<b>15.00 mL (Equivalence)</b>	<b>5.91</b>	25.00 mL	1.40
14.00 mL	9.49	15.01 mL	4.35	30.00 mL	1.30
14.20 mL	9.39	15.05 mL	3.65	35.00 mL	1.21

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



**Example 4:** 10.00 mL of  $\text{NH}_3(aq)$  is titrated with 0.0650 M of  $\text{HClO}_4(aq)$ .

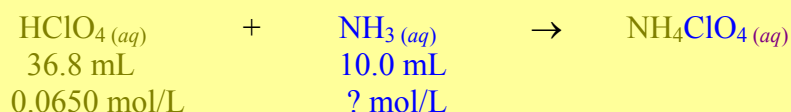
- Write the net ionic equation.
- Calculate the initial concentration of  $\text{NH}_3(aq)$  if the equivalence point is reached at 36.8 mL of acid added.

**a. Chemical Species Present:**  $\text{H}_3\text{O}^+$ ,  $\text{ClO}_4^-$ ,  $\text{NH}_3$ ,  $\text{H}_2\text{O}$  (Do NOT break up Weak Bases)  
 SA SB A/B (SB – Strongest Base from the list)

**Net-Ionic Equation:**  $\text{H}_3\text{O}^+(aq) + \text{NH}_3(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NH}_4^+(aq)$



b. For Acid-Base Stoichiometry, do NOT write the Net-Ionic Equation. Write the Molecular Equation to do Stoichiometry. Strong base will drive the weak acid towards completion (one way – not equilibrium) at equivalence point.



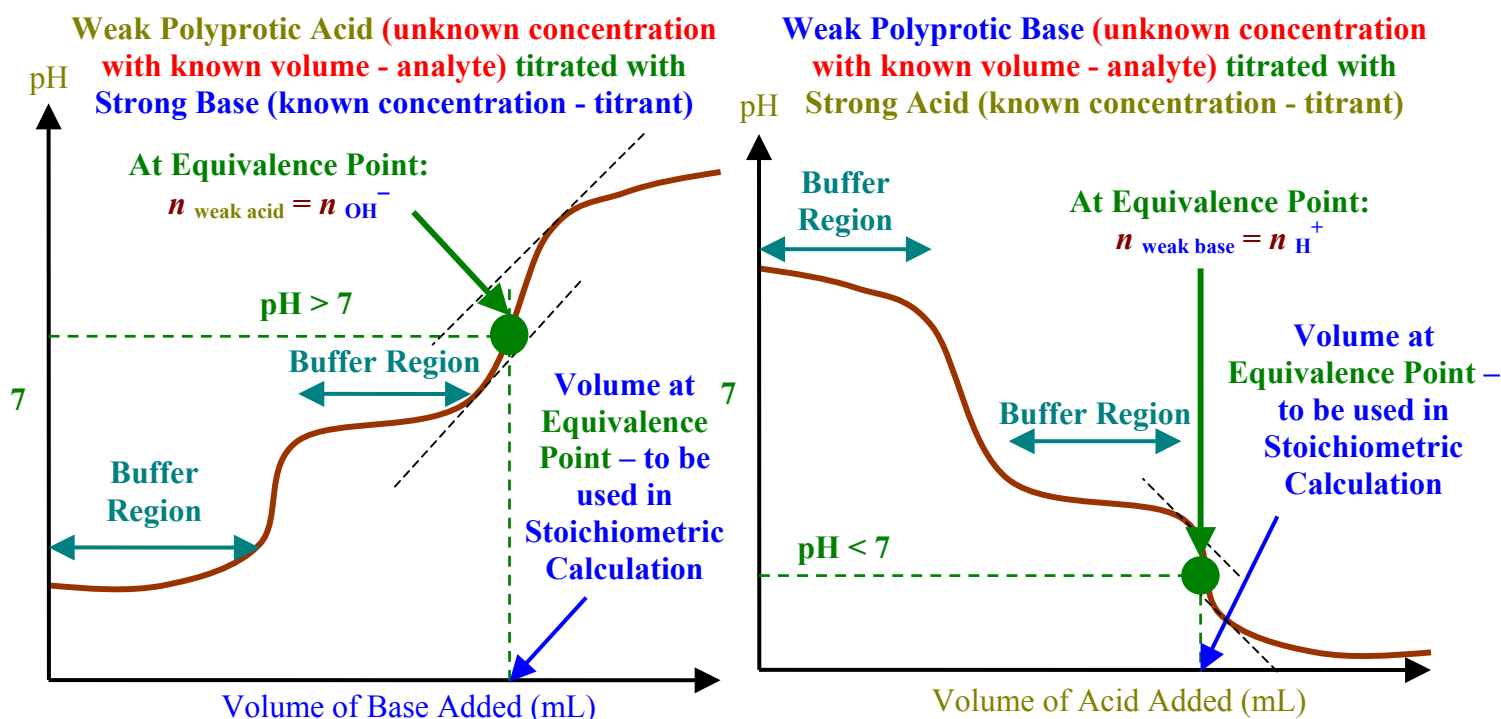
$$\textcircled{1} n_{\text{HClO}_4} = CV = (0.0650 \text{ mol/L})(36.8 \text{ mL}) = 2.392 \text{ mmol}$$

$$\textcircled{2} n_{\text{NH}_3} = 2.392 \text{ mmol HClO}_4 \times \frac{1 \text{ mol NH}_3}{1 \text{ mol HClO}_4} = 2.392 \text{ mmol NH}_3$$

$$\textcircled{3} [\text{NH}_3] = \frac{n}{V} = \frac{2.392 \text{ mmol}}{10.0 \text{ mL}}$$

$$[\text{NH}_3] = 0.239 \text{ mol/L} = 239 \text{ mmol/L}$$

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



**Example 5:** 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.

We use the **second equivalence point** because  $\text{Na}_2\text{S}$  can accept two protons.



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

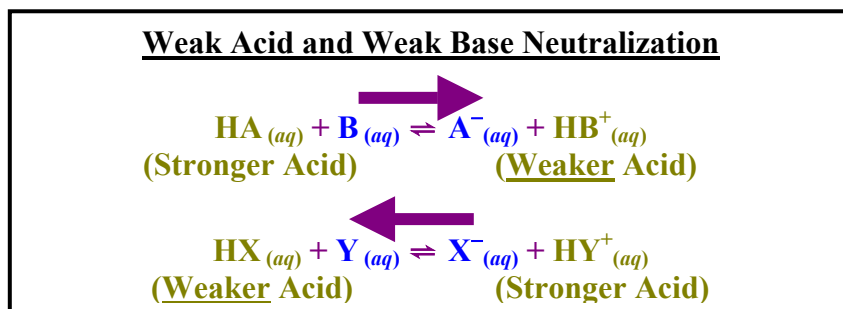
$$\textcircled{2} n_{\text{Na}_2\text{S}} = 0.9758 \text{ mmol HBr} \times \frac{1 \text{ mol Na}_2\text{S}}{2 \text{ mol HBr}} = 0.4879 \text{ mmol Na}_2\text{S}$$

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

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

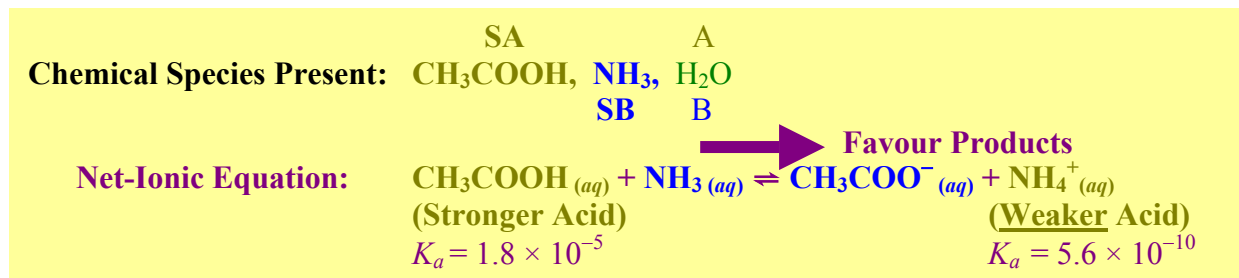
**1. The reaction is NOT one way, but in equilibrium.**

- When writing out the equilibrium system, **list all major species first (Do NOT Break up any Weak Acids and Weak Bases)**. Then, **reacts the Strongest Acid (SA – donates a proton) and the Strongest Base (SB – accepts a proton) from the list.**
- Using the Table of the Relative Strengths of Acids, decide on **which side of the equilibrium is the system favoured.** (The equilibrium favours the side **with the Weakest Acid (Lower  $K_a$ )** – because the stronger acid of the other side dissociates more.)

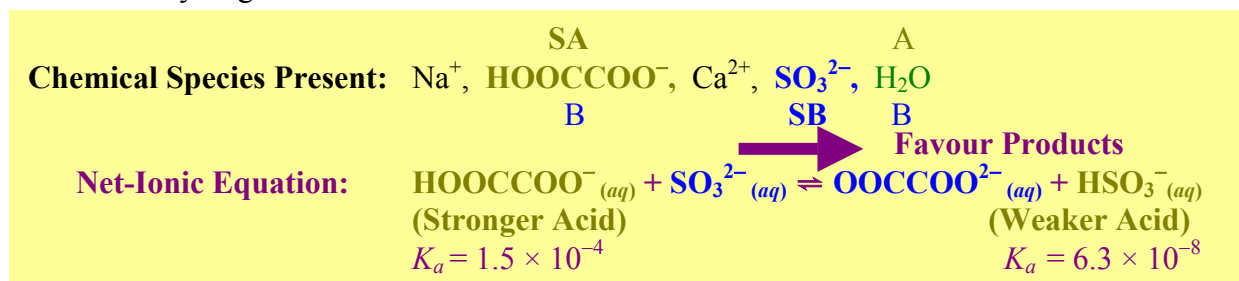


**Example 6:** Write the net-ionic equation for the following neutralizations. Indicate the side of the equilibrium system that is favoured.

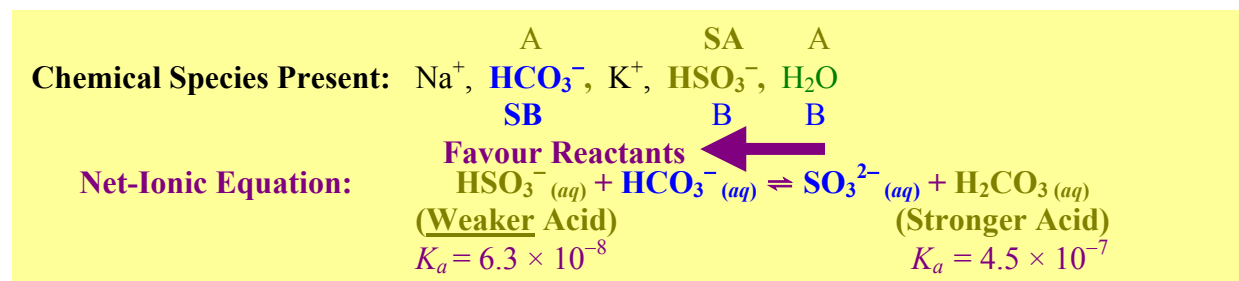
a. Acetic Acid reacts with Ammonia



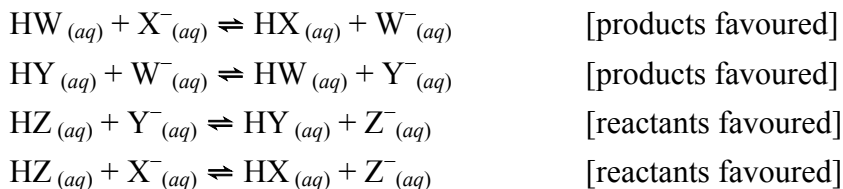
b. Sodium Hydrogen Oxalate reacts with Calcium Sulfite



c. Sodium Hydrogen Carbonate reacts with Potassium Hydrogen Sulfite

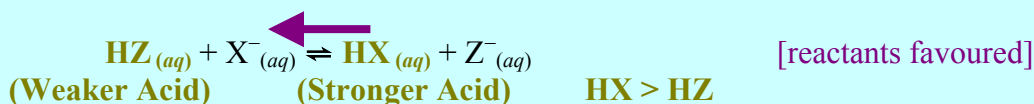
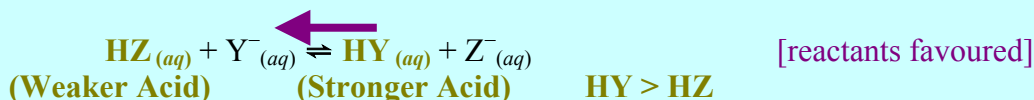
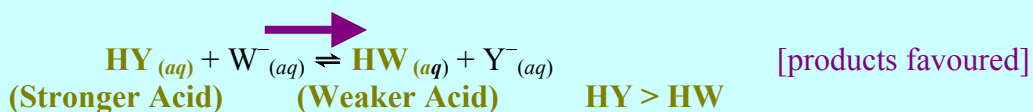
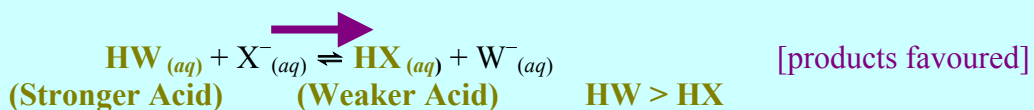


**Example 7:** Four different weak acids are reacted with various conjugate bases, and the results are recorded as follow.



List the acids in order of **increasing** strength.

First, we have to decide which acid is the weaker acid for each of these reactions.



Putting the inequalities together, we obtain in order of increasing acid strengths:



### Assignment

21.2 pg. 630 #23 and 24; pg. 640 #46 to 51;  
pg. 641 #59, 61, 65, 66 and 70

### Assignment

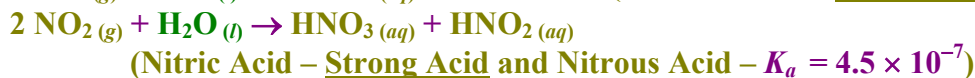
Chapter 21 Review pg. 642 #73, 75 and 77;  
pg. 643 #1 to 3, 5, 10 to 12

### 21.3: Environmental Issues Involving Acids and Bases

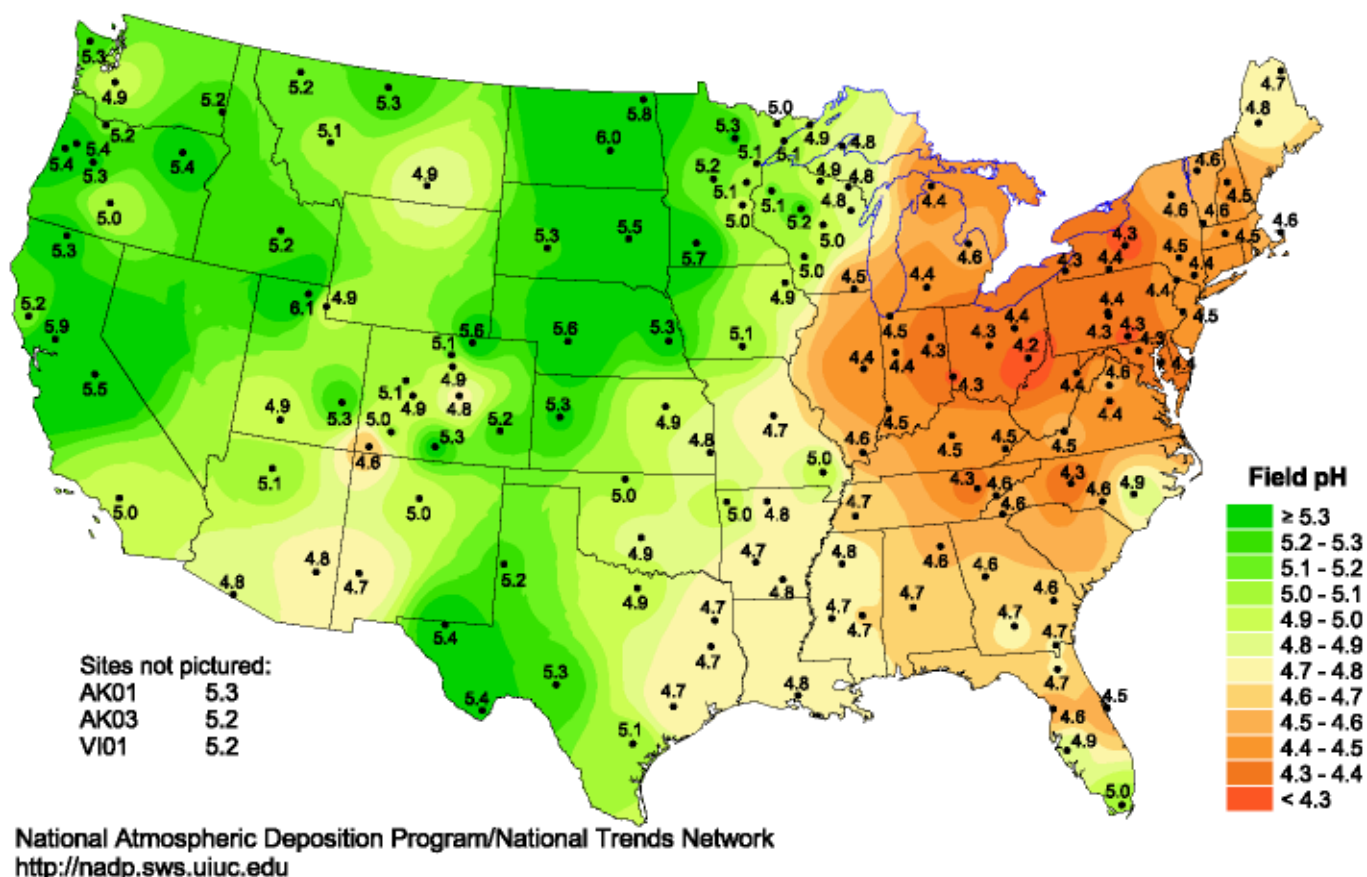
**Acid Rain:** - precipitation of water (rain) or (snow) that has a low pH (less than 5.0).  
 - natural rain is slightly acidic due to the small amount of  $\text{CO}_2$  in the atmosphere.



- however, the burning of fossil fuel by vehicles and industrial activities, and also emits other oxides such as  $\text{NO}_2(g)$ ,  $\text{SO}_2(g)$  and  $\text{SO}_3(aq)$ . These gaseous oxides (Lewis acids) combined with water in the atmosphere (Lewis base) to form various strong and weak acids. Thus, lowering the pH of rain and snow.

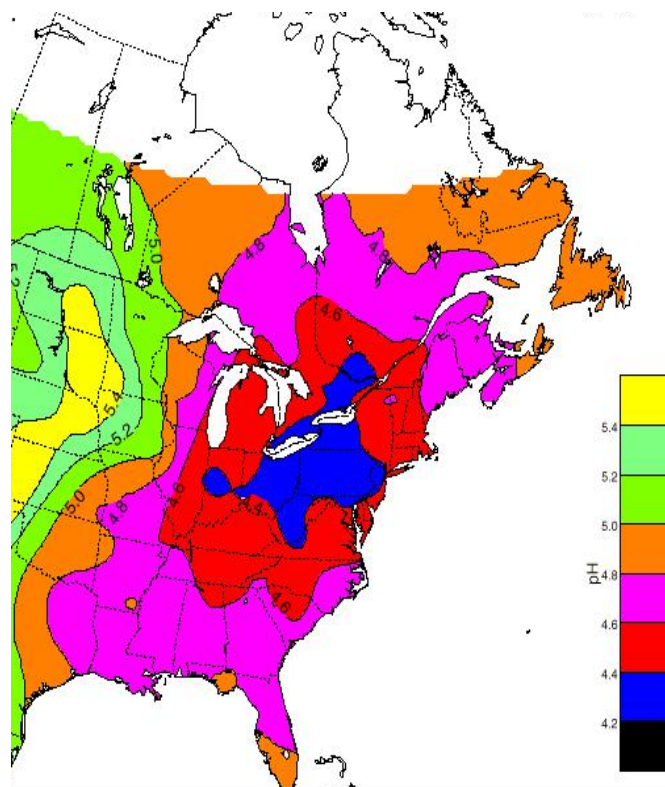
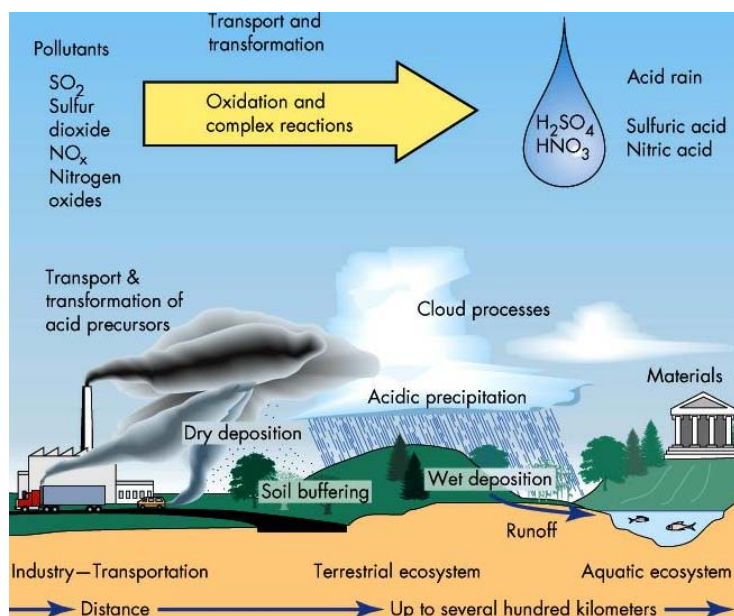


#### Hydrogen ion concentration as pH from measurements made at the field laboratories, 1999



Measurements of pH in the continental U.S in 1999. Note the most severely affected regions are concentrated in the East Coast as well as in the major cities.





(Above) The Formation of Acid Rain.

(Right) Measurements of pH in Eastern Canada and U.S. in 1996. The lowest pHs are in the Great Lakes Region where most of the heavy industries are located.

### Effects of Acid Rain:

1. **Marine Life (especially in-land lakes and rivers) is affected greatly.** Acid Rain lowers the pH that can destroy all life (plants and animals). Usually, acid lake and river appears to be crystal clear, with no sign of algae, plant and animal life.
2. **Vegetation (including forests) can be decimated.** Since the **soil acidity level are lowered**, further plant growth may not be possible and the surrounding **ecosystem is destroyed** as a result. Forests affected by acid rain appear to be charred with trees having no leaves and all branches are brittle.



(Left) an Acid Lake. Note the clear water with the absence of life. (Right) an Acid Forest. Note the lack of leaves, the brittle branches, and the charred looked of the tree trunks.

3. In areas of where there is a sustain amount of humidity, acid rain can turn into **Acid Fog or Smog ( $\text{SO}_2 + \text{H}_2\text{O}$ ) – Combination of Smoke from Coal Burning and Fog**. **Humans may experience severe respiratory problems as a result**. The most notable case was in London, England and surrounding areas in December 1952 when smog was so serious that there were estimated 4700 deaths.
4. **Acid deposition** (acid rain depositing onto a solid – metal or base that causes corrosion and neutralization) on city structures such as historical statues (usually made of limestone,  $\text{CaCO}_3$ ) and buildings can cause expensive repairs and lost of priceless arts.



(Above) Smog in London during December, 1952.  
(Below) Smog over the Los Angeles skyline.  
(Right) Acid Depositions on Statues at the Schloss Sanssouci Palace in Berlin, Germany. Note the black “burn” deposits on the statues.



#### Ways to Reduce Acid Rain:

1. **Burn Less Fossil Fuel:** drive more fuel efficient vehicles, develop alternate ways to generate electricity without combustion of fossil fuel, and reduce the amount of electricity used are all possible ways to cut down on gaseous oxides emissions.
2. **Enact Legislation to Regulate Emissions:** governments can pass legislation on industry to regulate the amount of gases that cause acid rain. Better filtering technology and frequent inspections of industrial facilities can hold companies accountable to these environmental laws.
3. **Adding Limestone** to acid lakes and rivers can raise the pH and reintroduce marine life that has been affected by acid rain.