Unit 8: ACIDS AND BASES

Chapter 15: Acids and Bases

15.1: What are Acids and Bases?
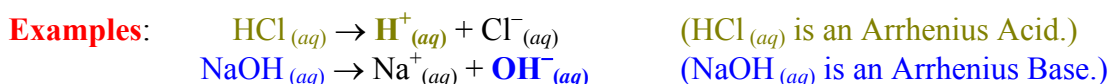
Physical and Chemical Properties of Acid and Base

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).
Conducts Electricity (Strong or Weak electrolytes)	Conducts Electricity (Strong or Weak electrolytes)
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

pH Scale

$$\overset{\displaystyle 0 \qquad\qquad \textbf{Acidic} \qquad\qquad 7 \qquad\qquad \textbf{Basic} \qquad\qquad 14}{\xleftrightarrow{\hspace{10cm}}}$$
Neutral

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**.
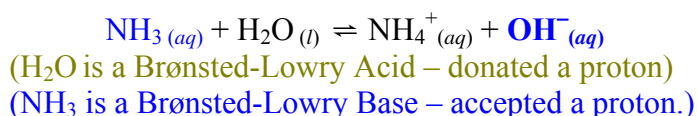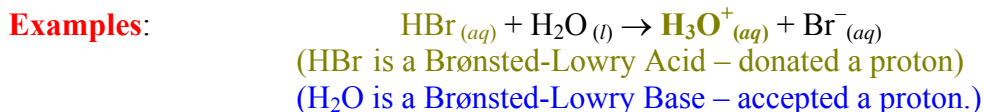
    Examples: $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$ ($HCl_{(aq)}$ is an Arrhenius Acid.)
                           $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$ ($NaOH_{(aq)}$ is an Arrhenius Base.)

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.

    Examples: $HBr_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Br^-_{(aq)}$
                        (HBr is a Brønsted-Lowry Acid – donated a proton)
                        (H_2O is a Brønsted-Lowry Base – accepted a proton.)

                     $NH_{3\ (aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+_{(aq)} + OH^-_{(aq)}$
                        (H_2O is a Brønsted-Lowry Acid – donated a proton)
                        (NH_3 is a Brønsted-Lowry Base – accepted a proton.)

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

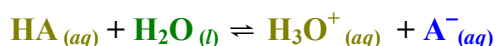**Conjugate Acid-Base Pair**: - the (**acid**/**conjugate base**) or (**base**/**conjugate acid**) pairs.

$$HA_{(aq)} \quad + \quad H_2O_{(l)} \quad \rightleftharpoons \quad H_3O^+_{(aq)} \quad + \quad A^-_{(aq)}$$
(Acid)   **(Base)**   **(Conjugate Acid)**   **(Conjugate Base)**

Conjugate Acid-Base Pairs: HA/A⁻ and H₂O/H₃O⁺

$$B_{(aq)} \quad + \quad H_2O_{(l)} \quad \rightleftharpoons \quad HB^+_{(aq)} \quad + \quad OH^-_{(aq)}$$
(Base)   **(Acid)**   **(Conjugate Acid)**   **(Conjugate Base)**

Conjugate Acid-Base Pairs: B/HB⁺ and H₂O/OH⁻

Brønsted-Lowry Acid Dissociation	**Brønsted-Lowry Base Dissociation**
$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$	$B_{(aq)} + H_2O_{(l)} \rightleftharpoons HB^+_{(aq)} + OH^-_{(aq)}$

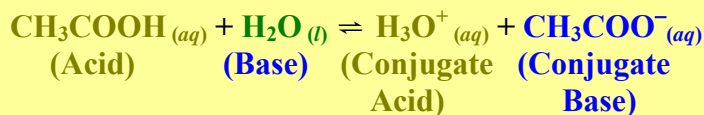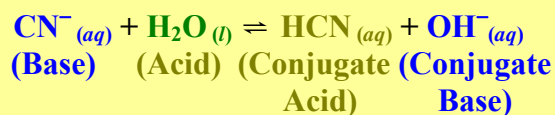**Example 1**: Write the Brønsted-Lowry dissociation reaction of the following. Identify the Brønsted-Lowry acid and base, along with the conjugate acid and base. Determine the conjugate acid-base pairs.

a. Acetic acid ($CH_3COOH_{(aq)}$)

b. Cyanide Ion ($CN^-_{(aq)}$)

$$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$
(Acid)   **(Base)**   **(Conjugate   (Conjugate**
                              Acid)   Base)

Conjugate Acid-Base Pairs:
CH₃COOH/CH₃COO⁻ and H₂O/H₃O⁺

$$CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$$
(Base)   **(Acid)**   **(Conjugate   (Conjugate**
                          Acid)   Base)

Conjugate Acid-Base Pairs:
CN⁻/HCN and H₂O/OH⁻

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 2: Name the following acids.

a. HI $_{(aq)}$

hydrogen iod*ide*
*hydro*iod*ic acid*

b. $HNO_{3\,(aq)}$

hydrogen nitr*ate*
nitr*ic acid*

c. $HNO_{2\,(aq)}$

hydrogen nitr*ite*
nitr*ous acid*

Example 3: Provide chemical formulas for the following acids.

a. hydrocyanic acid   b.   oxalic acid   c. sulfurous acid

*hydro*cyan*ic acid*

hydrogen cyan*ide* $\Rightarrow$ H$^+$ & CN$^-$

HCN $_{(aq)}$

oxal*ic acid* $\Rightarrow$ **hydrogen oxal*ate***

H$^+$ & OOCCOO^{2-}

H$_2$OOCCOO $_{(aq)}$

sulfur*ous acid*

$\Rightarrow$ **hydrogen sulf*ite***
H$^+$ & SO$_3$$^{2-}$

H$_2$SO$_3$ $_{(aq)}$

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 OH$^-$ ion.
- the naming of these ionic bases are "metal name" hydroxide.

Example 4: Name the following bases.

a.   KOH $_{(aq)}$

K$^+$ = potassium
OH$^-$ = hydroxide   **potassium
hydroxide**

b.   Ba(OH)$_2$ $_{(aq)}$

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

Example 5: Provide chemical formulas for the following acids.

a. strontium hydroxide

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

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

b.   rubidium hydroxide

Rb$^+$ and OH$^-$ $\Rightarrow$ **RbOH**

Need 1 Rb$^+$ & 1 (OH$^-$) to balance charges

Strengths of Acids and Bases

Strong Acids: - acids that <u>**dissociate completely (100%) in water**</u>.
- **the equilibrium position strongly <u>favours the products.</u>**
- at equilibrium, the original acid concentration, **[HA]$_{eq}$ ≈ 0; [H$_3$O$^+$]$_{eq}$ = [A$^-$]$_{eq}$ = [HA]$_0$.**
- the <u>**conjugate base, A$^-$**</u>, of a strong acid, HA <u>**is itself a weak base**</u> (cannot easily accept
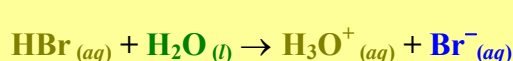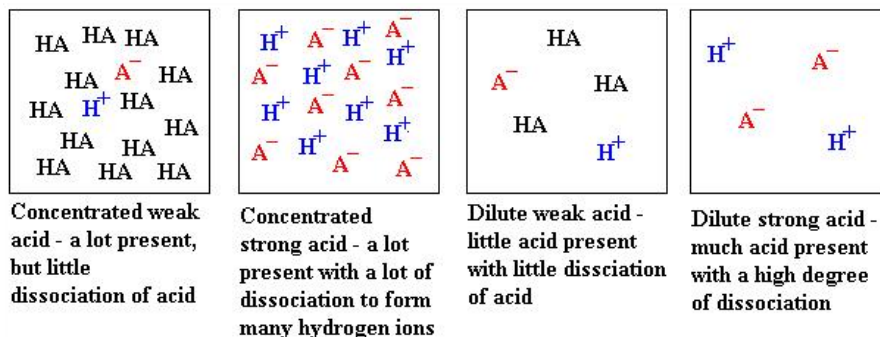protons to do the reverse reaction).
Note: **Strong Acids DO NOT MEAN that they are VERY CORROSIVE. It's [H$_3$O$^+$] that defines <u>acidity</u>.**

| | HA $_{(aq)}$ | + | H$_2$O $_{(l)}$ | $\rightleftharpoons$ | H$_3$O$^+$ $_{(aq)}$ | + | A$^-$ $_{(aq)}$ |
| | **(Strong Acid)** | | **(Base)** | | **(Conjugate Acid)** | | **(Conjugate Base-Weak)** |

	[HA]		[H$_3$O$^+$]	[A$^-$]
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$_2$SO$_4$ $_{(aq)}$ and HNO$_3$ $_{(aq)}$

Example 6: Write the dissociation reaction of 0.250 M of HBr $_{(aq)}$ and determine its [H$_3$O$^+$].

HBr $_{(aq)}$ + H$_2$O $_{(l)}$ $\rightarrow$ H$_3$O$^+$ $_{(aq)}$ + Br$^-$ $_{(aq)}$

(Strong Acid means [HBr]$_0$ = [H$_3$O$^+$])

[HBr]$_0$ = 0.250 M   **[H$_3$O$^+$] = 0.250 M**

Weak Acids: - acids that <u>**dissociate LESS than 100% in water**</u>.
            -**the equilibrium position strongly <u>favours the reactants.</u>**
            -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 <u>**conjugate base, A^-**</u>, of a weak acid **<u>is itself a stronger weak base</u>** (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**
      <u>**acidity**</u>. **At a high enough concentration, a weak acid can be corrosive.**

$$HA_{(aq)} \quad + \quad H_2O_{(l)} \quad \rightleftharpoons \quad H_3O^+_{(aq)} \quad + \quad A^-_{(aq)}$$
(Weak Acid)   **(Base)**   **(Conjugate Acid)** **(Conjugate Base-Stronger)**

	$[HA]$		$[H_3O^+]$	$[A^-]$
Initial	x		0	0
Change	$-y$ (where $y \ll x$)		$+y$	$+y$
Equilibrium	$(x-y) \approx x$		y	y

Examples: Some Weak Acids: $HOOCCOOH_{(aq)}$, $H_2SO_{3\,(aq)}$, $HSO_4^-{}_{(aq)}$, $H_3PO_{4\,(aq)}$, $HNO_{2\,(aq)}$, $H_3C_6H_5O_{7\,(aq)}$,
                 $HF_{(aq)}$, $HCOOH_{(aq)}$, $C_6H_8O_{6\,(aq)}$, $C_6H_5COOH_{(aq)}$, $CH_3COOH_{(aq)}$, $H_2CO_{3\,(aq)}$,
                 $H_2S_{(aq)}$, $HOCl_{(aq)}$, $HCN_{(aq)}$, $NH_4^+{}_{(aq)}$, and $H_3BO_{3\,(aq)}$
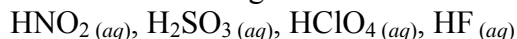

Concentrated weak acid - a lot present, but little dissociation of acid   Concentrated strong acid - a lot present with a lot of dissociation to form many hydrogen ions   Dilute weak acid - little acid present with little dissciation of acid   Dilute strong acid - much acid present with a high degree of dissociation

<u>Check out Strong and Weak Acids Animations at</u>
(<u>http://www.sgc.peachnet.edu/users/larnold/WWW/courses/1212/rev1212.html</u>)
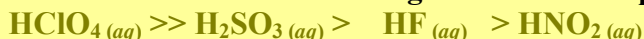
<u>Relative Strength of Acids and Conjugate Bases</u>:

The **<u>stronger the acid</u>**, the **<u>weaker its conjugate base</u>**. The **<u>weaker the acid</u>**, the **<u>stronger its conjugate base</u>**.

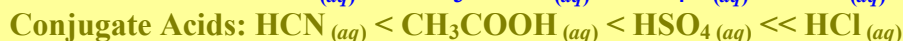**Example 7**: Order the following acids from the strongest to the weakest.
                $HNO_{2\,(aq)}$, $H_2SO_{3\,(aq)}$, $HClO_{4\,(aq)}$, $HF_{(aq)}$

> **According to the Acid and Bases Relative Strength Table on the previous page:**
> $HClO_{4\,(aq)} \gg H_2SO_{3\,(aq)} > \quad HF_{(aq)} \quad > HNO_{2\,(aq)}$

Example 8: Order the following conjugate base from the strongest to the weakest.
                $CN^-{}_{(aq)}$, $SO_4^{2-}{}_{(aq)}$, $Cl^-{}_{(aq)}$, $CH_3COO^-{}_{(aq)}$

> **According to the Acids and Bases Relative Strength Table on the previous page:**
>   **Bases:** $CN^-{}_{(aq)} > CH_3COO^-{}_{(aq)} > SO_4^{2-}{}_{(aq)} \gg Cl^-{}_{(aq)}$
> **Conjugate Acids:** $HCN_{(aq)} < CH_3COOH_{(aq)} < HSO_{4\,(aq)} \ll HCl_{(aq)}$

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

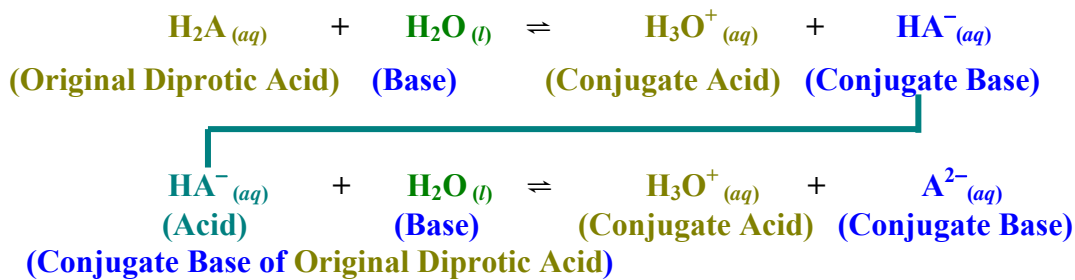| Acid Name	Acid Formula	Conjugate Base Formula	K_a
perchloric acid	$HClO_{4\,(aq)}$	$ClO_4^{-}{}_{(aq)}$	Very Large
hydroiodic acid	$HI_{(aq)}$	$I^-{}_{(aq)}$	Very Large
hydrobromic acid	$HBr_{(aq)}$	$Br^-{}_{(aq)}$	Very Large
hydrochloric acid	$HCl_{(aq)}$	$Cl^-{}_{(aq)}$	Very Large
sulfuric acid	$H_2SO_{4\,(aq)}$	$HSO_4^{-}{}_{(aq)}$	Very Large
nitric acid	$HNO_{3\,(aq)}$	$NO_3^{-}{}_{(aq)}$	Very Large
hydronium ion	$H_3O^+{}_{(aq)}$	$H_2O_{(l)}$	1
oxalic acid	$HOOCCOOH_{(aq)}$	$HOOCCOO^-{}_{(aq)}$	6.5×10^{-2}
sulfurous acid	$H_2SO_{3\,(aq)}$	HSO_3^-	1.5×10^{-2}
hydrogen sulfate ion	$HSO_4^{-}{}_{(aq)}$	SO_4^{2-}	1.2×10^{-2}
chlorous acid	$HClO_{2\,(aq)}$	$ClO_2^{-}{}_{(aq)}$	1.2×10^{-2}
phosphoric acid	$H_3PO_{4\,(aq)}$	$H_2PO_4^{-}{}_{(aq)}$	7.5×10^{-3}
arsenic acid	$H_3AsO_{4\,(aq)}$	$H_2AsO_4^{-}{}_{(aq)}$	5×10^{-3}
monochloracetic acid	$HC_2H_2ClO_2$	$C_2H_2ClO_2^{-}{}_{(aq)}$	1.35×10^{-3}
citric acid	$H_3C_6H_5O_{7\,(aq)}$	$H_2C_6H_5O_7^{-}{}_{(aq)}$	8.4×10^{-4}
hydrofluoric acid	$HF_{(aq)}$	$F^-{}_{(aq)}$	7.2×10^{-4}
nitrous acid	$HNO_{2\,(aq)}$	$NO_2^{-}{}_{(aq)}$	4.0×10^{-4}
methanoic (formic) acid	$HCOOH_{(aq)}$	$HCOO^-{}_{(aq)}$	1.8×10^{-4}
lactic acid	$HC_3H_5O_{3\,(aq)}$	$C_3H_5O_3^{-}{}_{(aq)}$	1.38×10^{-4}
ascorbic acid (vitamin C)	$H_2C_6H_6O_{6\,(aq)}$	$HC_6H_6O_6^{-}{}_{(aq)}$	7.9×10^{-5}
benzoic acid	$C_6H_5COOH_{(aq)}$	$C_6H_5COO^-{}_{(aq)}$	6.4×10^{-5}
hydrogen oxalate ion	$HOOCCOO^-{}_{(aq)}$	$OOCCOO^{2-}{}_{(aq)}$	6.1×10^{-5}
ethanoic (acetic) acid	$CH_3COOH_{(aq)}$	$CH_3COO^-{}_{(aq)}$	1.8×10^{-5}
dihydrogen citrate ion	$H_2C_6H_5O_7^{-}{}_{(aq)}$	$HC_6H_5O_7^{2-}{}_{(aq)}$	1.8×10^{-5}
hydrated aluminum (III) ion	$[Al(H_2O)_6]^{3+}$	$[AlOH(H_2O)_5]^{2+}$	1.4×10^{-5}
propanoic acid	$C_2H_5COOH_{(aq)}$	$C_2H_5COO^-{}_{(aq)}$	1.3×10^{-5}
hydrogen citrate ion	$HC_6H_5O_7^{2-}{}_{(aq)}$	$C_6H_5O_7^{3-}{}_{(aq)}$	4.0×10^{-6}
carbonic acid	$H_2CO_{3\,(aq)}$	$HCO_3^{-}{}_{(aq)}$	4.3×10^{-7}
hydrosulfuric acid	$H_2S_{(aq)}$	$HS^-{}_{(aq)}$	1.0×10^{-7}
hydrogen sulfite ion	$HSO_3^{-}{}_{(aq)}$	$SO_3^{2-}{}_{(aq)}$	1.0×10^{-7}
dihydrogen arsenate ion	$H_2AsO_4^{-}{}_{(aq)}$	$HAsO_4^{2-}{}_{(aq)}$	8×10^{-8}
dihydrogen phosphate ion	$H_2PO_4^{-}{}_{(aq)}$	$HPO_4^{2-}{}_{(aq)}$	6.2×10^{-8}
hypochlorous acid	$HOCl_{(aq)}$	$OCl^-{}_{(aq)}$	3.5×10^{-8}
hypobromous acid	$HOBr_{(aq)}$	$OBr^-{}_{(aq)}$	2×10^{-9}
hydrocyanic acid	$HCN_{(aq)}$	$CN^-{}_{(aq)}$	6.2×10^{-10}
hydrogen arsenate ion	$HAsO_4^{2-}{}_{(aq)}$	$AsO_4^{3-}{}_{(aq)}$	6×10^{-10}
boric acid	$H_3BO_{3\,(aq)}$	$H_2BO_3^{-}{}_{(aq)}$	5.8×10^{-10}
ammonium ion	$NH_4^{+}{}_{(aq)}$	$NH_{3\,(aq)}$	5.6×10^{-10}
Phenol	$C_6H_5OH_{(aq)}$	$C_6H_5O^-{}_{(aq)}$	1.6×10^{-10}
hydrogen carbonate ion	$HCO_3^{-}{}_{(aq)}$	$CO_3^{2-}{}_{(aq)}$	5.6×10^{-11}
hypoiodous acid	$HOI_{(aq)}$	$OI^-{}_{(aq)}$	2×10^{-11}
hydrogen ascorbate ion	$HC_6H_6O_6^{-}{}_{(aq)}$	$C_6H_6O_6^{2-}{}_{(aq)}$	1.6×10^{-12}
hydrogen phosphate ion	$HPO_4^{2-}{}_{(aq)}$	$PO_4^{3-}{}_{(aq)}$	4.8×10^{-13}
water (55.49 mol/L)	$H_2O_{(l)}$	$OH^-{}_{(aq)}$	$1.0 \times 10^{-14} = K_w$
hydrogen sulfide ion	$HS^-{}_{(aq)}$	$S^{2-}{}_{(aq)}$	$\sim 10^{-19}$

Increasing Acid Strength

Increasing Base Strength

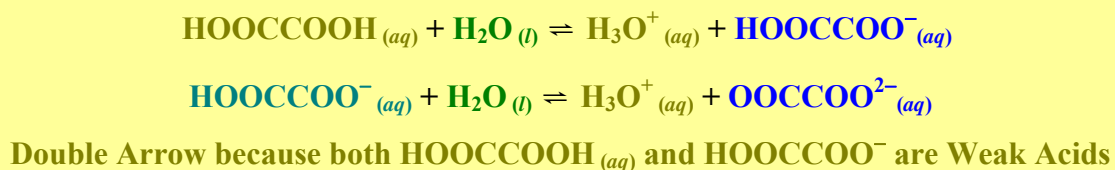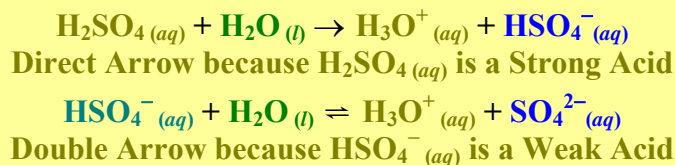**Monoprotic Acids**: - acids that can donate a maximum of one proton.

Example 9: Write the dissociation reaction for the following monoprotic acids.

a. $HI_{(aq)}$

$HI_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + I^-_{(aq)}$
Direct Arrow because HI$_{(aq)}$ is a Strong Acid

b. $HCOOH_{(aq)}$

$HCOOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HCOO^-_{(aq)}$
Double Arrow because HCOOH$_{(aq)}$ is a Weak Acid

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

$$H_2A_{(aq)} \quad + \quad H_2O_{(l)} \quad \rightleftharpoons \quad H_3O^+_{(aq)} \quad + \quad HA^-_{(aq)}$$
(Original Diprotic Acid)   (Base)   (Conjugate Acid)   (Conjugate Base)

$$HA^-_{(aq)} \quad + \quad H_2O_{(l)} \quad \rightleftharpoons \quad H_3O^+_{(aq)} \quad + \quad A^{2-}_{(aq)}$$
(Acid)   (Base)   (Conjugate Acid)   (Conjugate Base)
(Conjugate Base of Original Diprotic Acid)

Example 10: Write the stepwise dissociation reaction for the following diprotic acids.

a. $H_2SO_{4\ (aq)}$

$H_2SO_{4\ (aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + HSO_4^-_{(aq)}$
Direct Arrow because H$_2$SO$_{4\ (aq)}$ is a Strong Acid
$HSO_4^-_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + SO_4^{2-}_{(aq)}$
Double Arrow because HSO$_4^-$$_{(aq)}$ is a Weak Acid
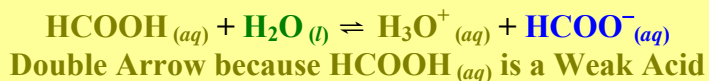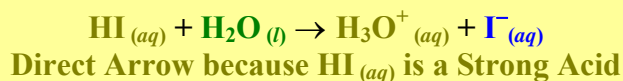
b. $HOOCCOOH_{(aq)}$

$HOOCCOOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HOOCCOO^-_{(aq)}$

$HOOCCOO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OOCCOO^{2-}_{(aq)}$

Double Arrow because both HOOCCOOH$_{(aq)}$ and HOOCCOO$^-$ are Weak Acids

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: $HOOCCOO^-_{(aq)}$, $HSO_4^-_{(aq)}$, $HSO_3^-_{(aq)}$, $HCO_3^-_{(aq)}$, $HS^-_{(aq)}$, $HC_6H_6O_6^-_{(aq)}$, and $H_2O_{(l)}$

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**.
- 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*.

Examples: Some Polyprotic Acids:

    a. Diprotic Acids: $H_2SO_{4\ (aq)}$, $HOOCCOOH_{(aq)}$, $H_2SO_{3\ (aq)}$, $H_2C_6H_6O_{6\ (aq)}$, $H_2CO_{3\ (aq)}$, $H_2S_{(aq)}$

       Diprotic Amphoteric Intermediates: $HSO_4^-{}_{(aq)}$, $HOOCCOO^-{}_{(aq)}$, $HSO_3^-{}_{(aq)}$, $HC_6H_6O_6^-{}_{(aq)}$, $HCO_3^-{}_{(aq)}$, $HS^-{}_{(aq)}$

    b. Triprotic Acids: $H_3PO_{4\ (aq)}$, $H_3AsO_{4\ (aq)}$, $H_3C_6H_5O_{7\ (aq)}$, $H_3BO_{3\ (aq)}$

       Triprotic Amphoteric Intermediates: $H_2PO_4^-{}_{(aq)}$, $HPO_4^{2-}{}_{(aq)}$, $H_2AsO_4^-{}_{(aq)}$, $HAsO_4^{2-}{}_{(aq)}$, $H_2C_6H_5O_7^-{}_{(aq)}$, $HC_6H_5O_7^{2-}{}_{(aq)}$, $H_2BO_3^-{}_{(aq)}$, $HBO_3^{2-}{}_{(aq)}$

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 ($H_2SO_{4\ (aq)}$), successively donate protons to finally reach $SO_4^{2-}{}_{(aq)}$
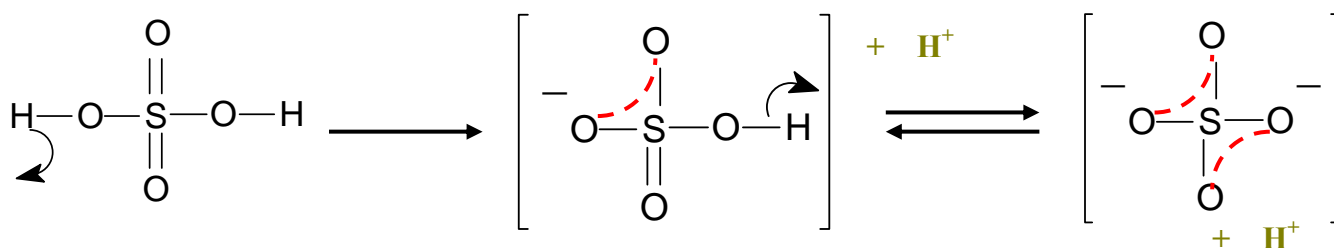
Organic Acids: - acids of organic compounds, commonly known as carboxylic acids.
        - alcohols can also be considered as organic acids, but they are really weak. This is because
          carboxylic acids have double-bonded oxygen for resonance to occur, whereas alcohols
          only contain single-bonded oxygen.

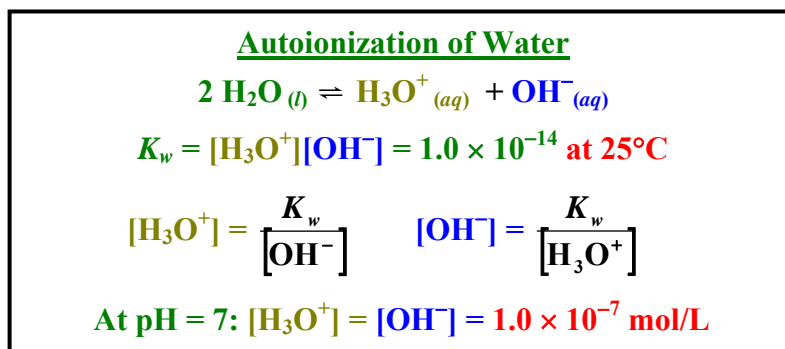**Example**: Ethanoic (Acetic) Acid is compared to Phenol

<div style="text-align:center">
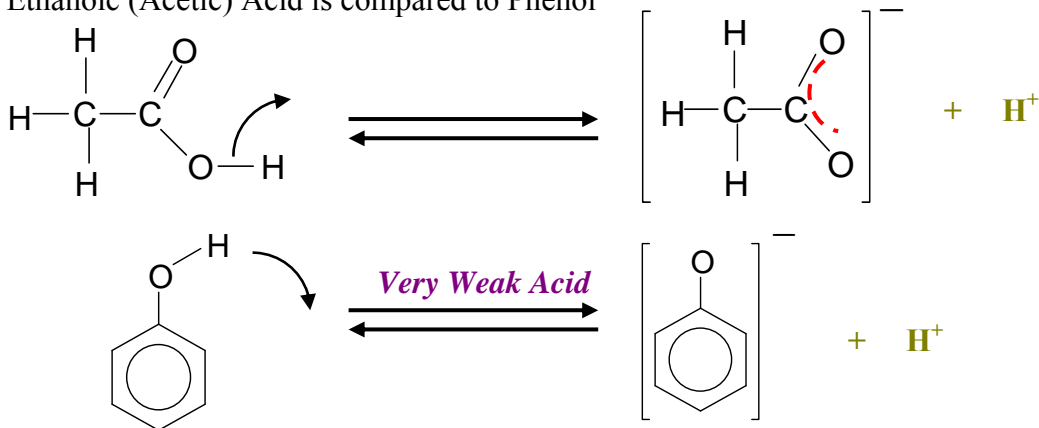
Assignment
15.1 pg. 538 #1 to 13

</div>

15.2: Acidity, Basicity and pH

Autoionization of Water: - the process where water dissociates into hydronium and hydroxide ions.
- water's dissociation constant is called **Self-Ionization Constant** ($K_w = 1.0 \times 10^{-14}$)
- when the solution is **neutral (pH = 7)**, $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/L

<div style="border:1px solid black;">

Autoionization of Water

$$2 \; H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25°C$$

$$[H_3O^+] = \frac{K_w}{[OH^-]} \qquad [OH^-] = \frac{K_w}{[H_3O^+]}$$

At pH = 7: $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/L

</div>

Example 1: Determine the $[H_3O^+]$ and/or $[OH^-]$ concentrations of the following solutions at 25°C.

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

$$K_w = [H_3O^+][OH^-]$$
$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}}$$
$$[H_3O^+] = 1.0 \times 10^{-10} \text{ mol/L}$$
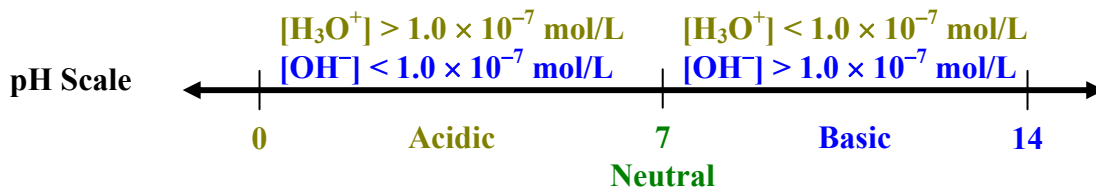
b. $[H_3O^+] = 5.0 \times 10^{-5}$ M

$$K_w = [H_3O^+][OH^-]$$
$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-5}}$$
$$[OH^-] = 2.0 \times 10^{-10} \text{ M}$$

pH Scale: - a logarithmic scale to **measure the acidity (relative $[H_3O^+]$)** of a solution.
- the **lower the pH, the more acidic** (less basic) is the solution (**more $[H_3O^+]$ and less $[OH^-]$**).
- the **higher the pH, the more basic** (less acidic) is the solution (**less $[H_3O^+]$ and more $[OH^-]$**).
- **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 ≠ 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 $[H_3O^+]$ by a factor of 10**; an increase of **2 on a pH scale means a decrease of $[H_3O^+]$ by a factor of 100.**

$[H_3O^+] > 1.0 \times 10^{-7}$ mol/L $[H_3O^+] < 1.0 \times 10^{-7}$ mol/L
$[OH^-] < 1.0 \times 10^{-7}$ mol/L $[OH^-] > 1.0 \times 10^{-7}$ mol/L

pH Scale ←————————————————————————→

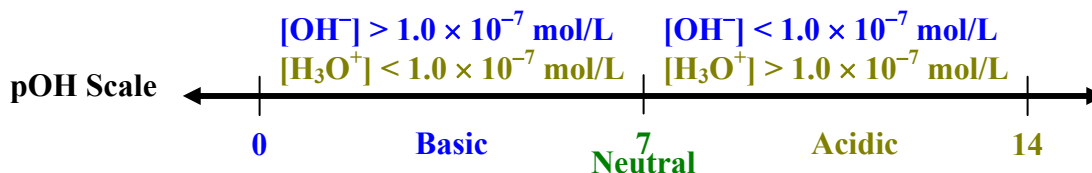0   **Acidic**   7   **Basic**   14

Neutral

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 [OH⁻])** of a solution.
- the **lower the pOH, the more basic** (less acidic) is the solution (**more [OH⁻]** and **less [H₃O⁺]**).
- the **higher the pOH, the less basic** (more acidic) is the solution (**less [OH⁻]** and **more [H₃O⁺]**).
- **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 ≠ 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 [OH⁻] by a factor of 10;** an **increase of 2 on a pOH scale means a decrease of [OH⁻] by a factor of 100.**

pOH Scale

$[OH^-] > 1.0 \times 10^{-7}$ mol/L $[OH^-] < 1.0 \times 10^{-7}$ mol/L
$[H_3O^+] < 1.0 \times 10^{-7}$ mol/L $[H_3O^+] > 1.0 \times 10^{-7}$ mol/L

0   **Basic**   **7**   **Acidic**   14
                              Neutral

---

pH and pOH Scales

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

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

$pH + pOH = 14.00$

Example 2: Calculate the $[H_3O^+]$ and the $[OH^-]$ for the following solutions.

a. pH = 4.00

$pH = -\log [H_3O^+]$
$[H_3O^+] = 10^{-pH} = 10^{-4.00}$

$\boxed{[H_3O^+] = 1.00 \times 10^{-4} \text{ M}}$

$pOH = 14 - pH$
$pOH = 14 - 4.00 = 10.00$

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

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

b. pOH = 3.00

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

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

$pH = 14 - pOH$
$pH = 14 - 3 = 11.00$

$pH = -\log [H_3O^+]$
$[H_3O^+] = 10^{-pH} = 10^{-11.00}$

$\boxed{[H_3O^+] = 1.00 \times 10^{-11} \text{ M}}$

c. pH = 12.83

$pH = -\log [H_3O^+]$
$[H_3O^+] = 10^{-pH} = 10^{-12.83}$

$\boxed{[H_3O^+] = 1.48 \times 10^{-13} \text{ M}}$

$K_w = [H_3O^+][OH^-]$

$[OH^-] = \dfrac{K_w}{[H_3O^+]} = \dfrac{1.0 \times 10^{-14}}{1.48 \times 10^{-13}}$

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

d. pOH = 9.67

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

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

$K_w = [H_3O^+][OH^-]$

$[H_3O^+] = \dfrac{K_w}{[OH^-]} = \dfrac{1.0 \times 10^{-14}}{2.14 \times 10^{-10}}$

$\boxed{[H_3O^+] = 4.68 \times 10^{-5} \text{ mol/L}}$

Calculating the pH of Strong Acid and 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 ($H_2O_{(l)}$) is always listed as a major species**.

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

$$HA_{(aq)} \quad + \quad H_2O_{(l)} \longrightarrow \quad H_3O^+_{(aq)} \quad + \quad A^-_{(aq)}$$
(Strong Acid)   **(Base)**   **(Conjugate Acid)**   **(Conjugate Base-Weak)**

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.**

$$MOH_{(aq)} \longrightarrow M^+_{(aq)} \quad + \quad OH^-_{(aq)}$$
(Strong Base)   **(Spectator Metal Cation)**   **(Major Species of Strong Base)**

Example 3: Determine the $[H_3O^+]$ and/or $[OH^-]$ concentrations of the following solutions at 25°C.
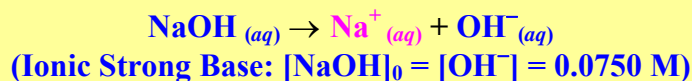
a. $[HCl] = 0.350$ mol/L

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$
(Strong Acid: $[HCl]_0 = [H_3O^+] = 0.350$ M)
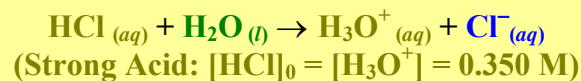
$$K_w = [H_3O^-][OH^-]$$
$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{0.350}$$
$$\boxed{[OH^-] = 2.86 \times 10^{-14} \text{ mol/L}}$$

b. $[NaOH] = 0.0750$ M

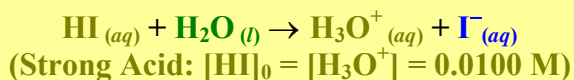$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$
(Ionic Strong Base: $[NaOH]_0 = [OH^-] = 0.0750$ M)

$$K_w = [H_3O^+][OH^-]$$
$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0\times10^{-14}}{0.0750}$$
$$\boxed{[H_3O^+] = 1.33 \times 10^{-13} \text{ mol/L}}$$

Example 4: Calculate the pH and the pOH for the following solutions.
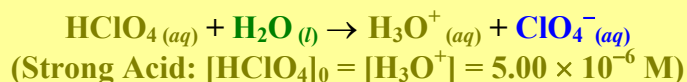
a. 0.0100 mol/L of $HI_{(aq)}$ (completely dissociates)

$$HI_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + I^-_{(aq)}$$
(Strong Acid: $[HI]_0 = [H_3O^+] = 0.0100$ M)

$pH = -\log[H_3O^+]$ $pOH = 14 - pH$
$pH = -\log(0.0100)$ $pOH = 14 - 2.00$
$\boxed{pH = 2.00}$ $\boxed{pOH = 12.00}$

b. 5.00×10^{-6} M of $HClO_{4\,(aq)}$ (completely dissociates)

$$HClO_{4\,(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + ClO_4^-_{(aq)}$$
(Strong Acid: $[HClO_4]_0 = [H_3O^+] = 5.00 \times 10^{-6}$ M)

$pH = -\log[H_3O^+]$ $pOH = 14 - pH$
$pH = -\log(5.00 \times 10^{-6})$ $pOH = 14 - 5.30$
$\boxed{pH = 5.30}$ $\boxed{pOH = 8.70}$

c. 1.50 M of $KOH_{(aq)}$

$$KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$$
(Ionic Strong Base: $[KOH]_0 = [OH^-] = 1.50$ M)

$pOH = -\log[OH^-]$ $pH = 14 - pOH$
$pOH = -\log(1.50)$ $pH = 14 - (-0.176)$
$\boxed{pOH = -0.176}$ $\boxed{pH = 14.176}$

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

$$Ba(OH)_{2\,(aq)} \rightarrow Ba^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$
(Ionic Strong Base:
$[OH^-] = 2[Ba(OH)_2] = 2.00 \times 10^{-3}$ M)
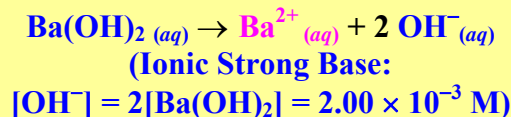
$pOH = -\log[OH^-]$ $pH = 14 - pOH$
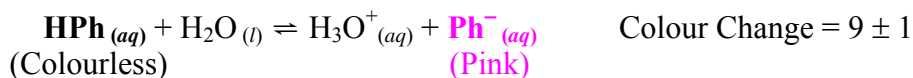$pOH = -\log(2.00 \times 10^{-3})$ $pH = 14 - (2.70)$
$\boxed{pOH = 2.70}$ $\boxed{pH = 11.30}$

<u>**Acid-Base Indicators**</u>: - 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.
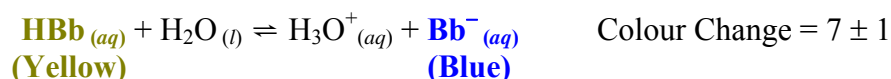- this is due to the acidic form of the indictor (HIn) has a different colour than its basic form (In⁻).

Example: Phenolphthalein (HPh $_{(aq)}$ /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)**.

$$\text{HPh }_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + \text{Ph}^-{}_{(aq)} \qquad \text{Colour Change} = 9 \pm 1$$
(Colourless) (Pink)

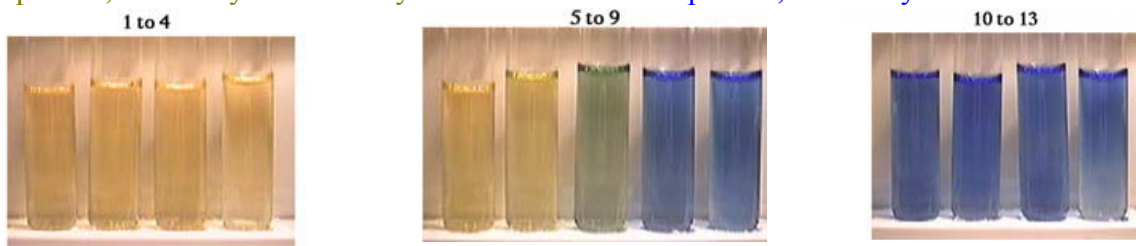At pH ≤ 9, phenolphthalein is colourless. At pH ≥ 10, phenolphthalein is pink.

Example: Bromothymol Blue (HBb $_{(aq)}$ /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)**.

$$\text{HBb }_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + \text{Bb}^-{}_{(aq)} \qquad \text{Colour Change} = 7 \pm 1$$
(Yellow) **(Blue)**

At pH ≤ 6, bromothymol blue is yellow. At pH ≥ 8, bromothymol blue is blue.

Example 5: A sample of rainwater is poured into five test tubes. A different indicator is added to each test-tube. Four of the observations are recorded in the table below.

methyl red	yellow
phenol red	yellow
bromocresol green	blue
phenolphthalein	colourless
bromothymol blue	?

Determine the pH of the rainwater and the predicted colour of the sample containing bromothymol blue.

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

Indicators	Colour Change	pH range	Colour Observed	Indicated pH
methyl red	red to yellow	4.8 to 6.0	Yellow	pH > 6.0
phenol red	yellow to red	6.6 to 8.0	Yellow	pH < 6.6
bromocresol green	yellow to blue	3.8 to 5.4	Blue	pH > 5.4
phenolphthalein	colourless to pink	8.2 to 10.0	colourless	pH < 8.2

From the indicated pH ranges, we can tell than the pH of the rain water is **between 6.0 to 6.6**. At such pH range, **bromothymol blue would appear green**. (The pH and colour range for bromothymol blue is 6.0 − 7.6; yellow to blue)

Acid-Base Indicators at 25°C (298.15 K)

Indicator	Suggested Abbreviation (s)	pH Range	Colour Change as pH Increases
methyl violet	$HMv_{(aq)}$ / $Mv^-_{(aq)}$	0.0 – 1.6	yellow to blue
cresol red	$H_2Cr_{(aq)}$ / $HCr^-_{(aq)}$	0.0 – 1.0	red to yellow
	$HCr_{(aq)}$ / $Cr^{2-}_{(aq)}$	7.0 – 8.8	yellow to red
thymol blue	$H_2Tb_{(aq)}$ / $HTb^-_{(aq)}$	1.2 – 2.8	red to yellow
	$HTb^-_{(aq)}$ / $Tb^{2-}_{(aq)}$	8.0 – 9.6	yellow to blue
orange IV	$HOr_{(aq)}$ / $Or^-_{(aq)}$	1.4 – 2.8	red to yellow
methyl orange	$HMo_{(aq)}$ / $Mo^-_{(aq)}$	3.2 – 4.4	red to yellow
bromocresol green	$HBg_{(aq)}$ / $Bg^-_{(aq)}$	3.8 – 5.4	yellow to blue
methyl red	$HMr_{(aq)}$ / $Mr^-_{(aq)}$	4.8 – 6.0	red to yellow
chlorophenol red	$HCh_{(aq)}$ / $Ch^-_{(aq)}$	5.2 – 6.8	yellow to red
bromothymol blue	$HBb_{(aq)}$ / $Bb^-_{(aq)}$	6.0 – 7.6	yellow to blue
phenol red	$HPr_{(aq)}$ / $Pr^-_{(aq)}$	6.6 – 8.0	yellow to red
phenolphthalein	$HPh_{(aq)}$ / $Ph^-_{(aq)}$	8.2 – 10.0	colourless to pink
thymolphthalein	$HTh_{(aq)}$ / $Th^-_{(aq)}$	9.4 – 10.6	colourless to blue
alizarin yellow R	$HAy_{(aq)}$ / $Ay^-_{(aq)}$	10.1 – 12.0	yellow to red
indigo carmine	$HIc_{(aq)}$ / $Ic^-_{(aq)}$	11.4 – 13.0	blue to yellow
1,3,5 - trinitrobenzene	$HNb_{(aq)}$ / $Nb^-_{(aq)}$	12.0 – 14.0	colourless to orange

Assignment
15.2 pg. 541 #1 to 5 (Practice); pg.544 #1 to 4 (Practice);
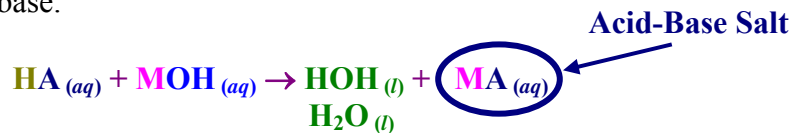**pg. 545 #1 to 4 (Practice); pg. 547 #1 to 15**

15.3: Neutralization and Titrations

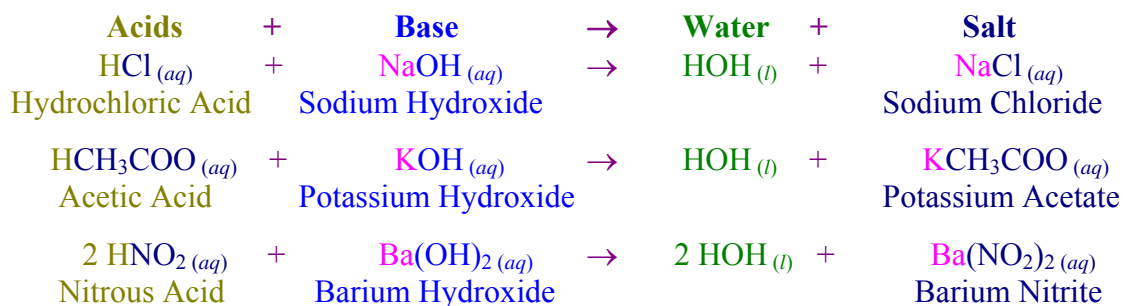**Neutralization**: - the reaction of an acid and a base to form water and an acid-base salt.

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.

Acid-Base Salt

General Neutralization Equation: $HA_{(aq)} + MOH_{(aq)} \rightarrow HOH_{(l)} + MA_{(aq)}$
                                                          $H_2O_{(l)}$

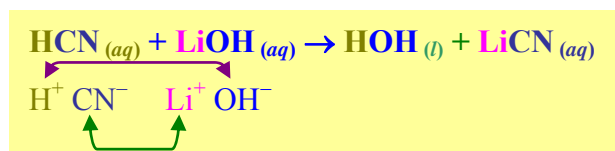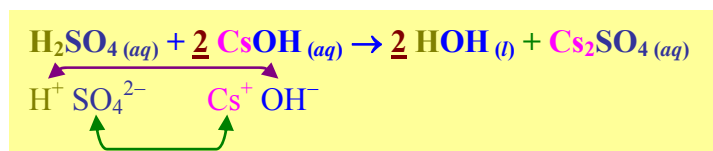| **Examples**:	**Acids**	+	**Base**	$\rightarrow$	**Water**	+	**Salt**
	$HCl_{(aq)}$	+	$NaOH_{(aq)}$	$\rightarrow$	$HOH_{(l)}$	+	$NaCl_{(aq)}$
	Hydrochloric Acid		Sodium Hydroxide				Sodium Chloride
	$HCH_3COO_{(aq)}$	+	$KOH_{(aq)}$	$\rightarrow$	$HOH_{(l)}$	+	$KCH_3COO_{(aq)}$
	Acetic Acid		Potassium Hydroxide				Potassium Acetate
	$2\ HNO_{2\ (aq)}$	+	$Ba(OH)_{2\ (aq)}$	$\rightarrow$	$2\ HOH_{(l)}$	+	$Ba(NO_2)_{2\ (aq)}$
	Nitrous Acid		Barium Hydroxide				Barium Nitrite

Example 1: Write a balanced chemical equation for the following neutralization reactions.

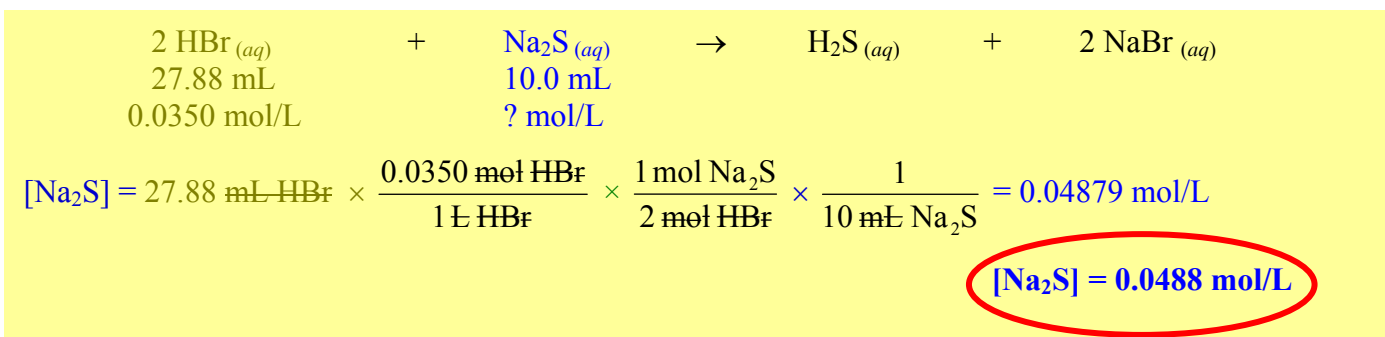a. hydrocyanic acid with lithium hydroxide   b. sulfuric acid with cesium hydroxide

$HCN_{(aq)} + LiOH_{(aq)} \rightarrow HOH_{(l)} + LiCN_{(aq)}$
$H^+\ CN^- \quad Li^+\ OH^-$

$H_2SO_{4\ (aq)} + 2\ CsOH_{(aq)} \rightarrow 2\ HOH_{(l)} + Cs_2SO_{4\ (aq)}$
$H^+\ SO_4^{2-} \quad Cs^+\ OH^-$

c. phosphoric acid with strontium hydroxide

$2\ H_3PO_{4\ (aq)} + 3\ Sr(OH)_{2\ (aq)} \rightarrow 6\ HOH_{(l)} + Sr_3(PO_4)_{2\ (s)}$
$H^+\ PO_4^{3-} \quad Sr^{2+}\ OH^-$
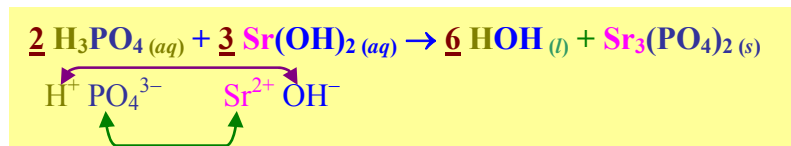
Example 2: Calculate the concentration of 10.00 mL of $Na_2S_{(aq)}$ (a base) reacts completely with 27.88 mL of 0.0350 M of hydrobromic acid ($HBr_{(aq)}$).

$2\ HBr_{(aq)}$	+	$Na_2S_{(aq)}$	$\rightarrow$	$H_2S_{(aq)}$	+	$2\ NaBr_{(aq)}$
27.88 mL		10.0 mL				
0.0350 mol/L		? mol/L				

$$[Na_2S] = 27.88\ \cancel{mL\ HBr} \times \frac{0.0350\ \cancel{mol\ HBr}}{1\ \cancel{L\ HBr}} \times \frac{1\ mol\ Na_2S}{2\ \cancel{mol\ HBr}} \times \frac{1}{10\ \cancel{mL}\ Na_2S} = 0.04879\ mol/L$$

$$[Na_2S] = 0.0488\ mol/L$$

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

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

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).

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

Titration Set-up
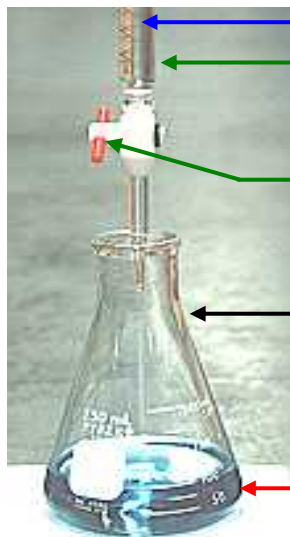
Titrant: - the solution of known concentration.

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.

pH Curve: - a graph that shows the how the pH changes as the titration proceeds (as the titrant is added to the analyzed).
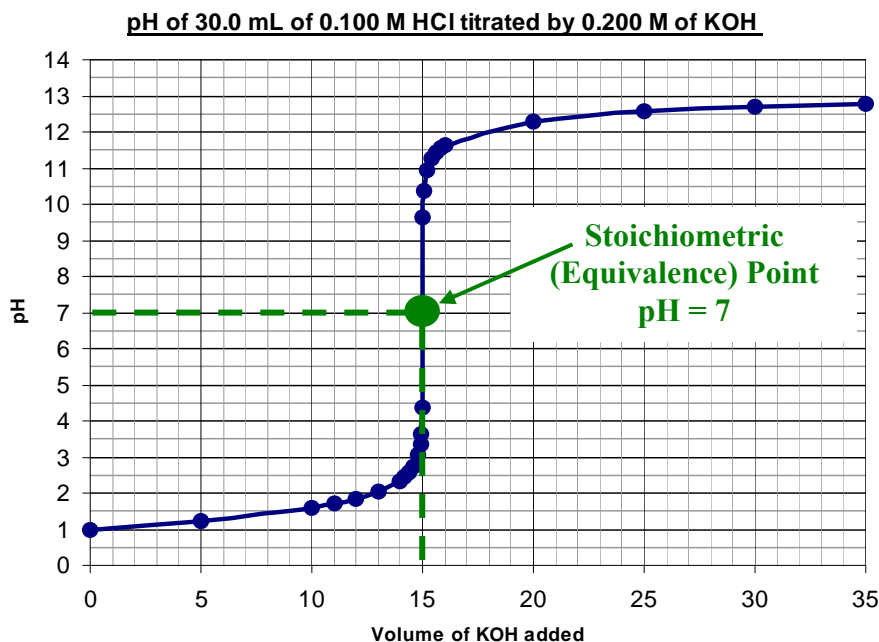
Inflection: - the part of the curve where there is a sudden rise or drop in pH.
- the midpoint of the inflection marks the equivalence point.

Titration Between Strong Acids and Strong Base: - **Equivalence Point** always occur at **pH = 7**.
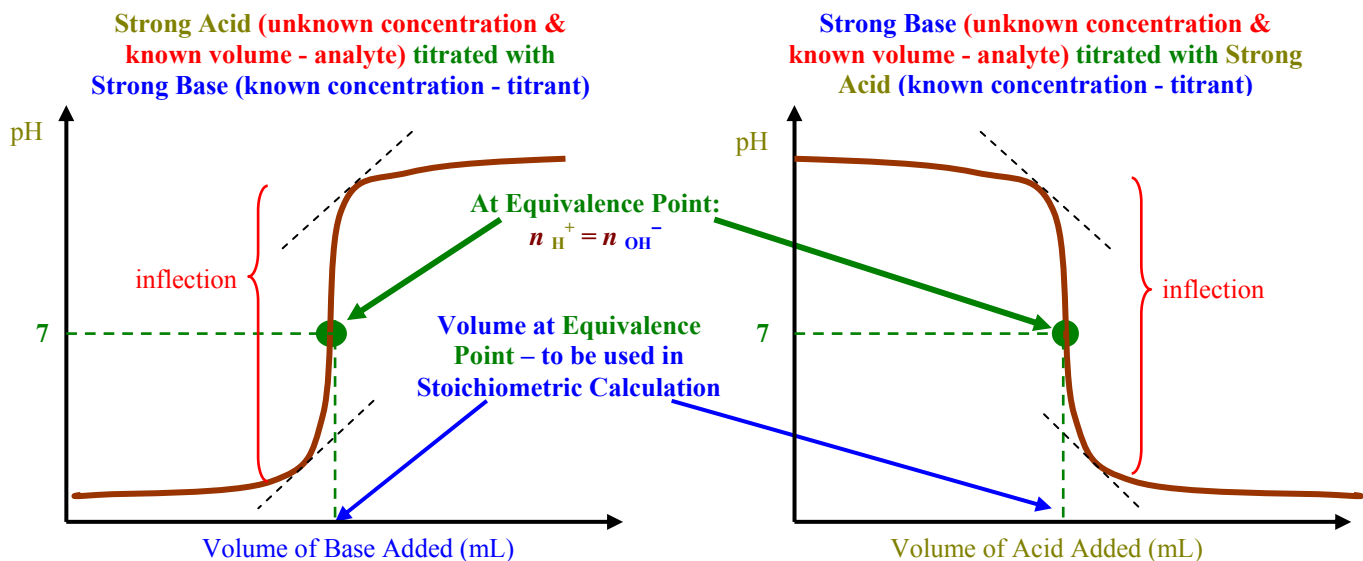
Example: **pH of 30.0 mL of 0.100 M HCl $_{(aq)}$ titrated by 0.200 M of KOH $_{(aq)}$**

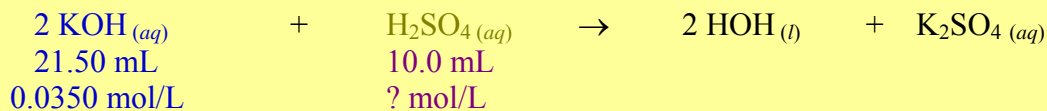| Volume of KOH added	pH	Volume of KOH added	pH	Volume of KOH added	pH
0.00 mL	1.00	14.60 mL	2.75	15.40 mL	11.25
5.00 mL	1.24	14.80 mL	3.05	15.60 mL	11.42
10.00 mL	1.60	14.90 mL	3.35	15.80 mL	11.54
11.00 mL	1.71	14.95 mL	3.65	16.00 mL	11.64
12.00 mL	1.85	14.99 mL	4.35	20.00 mL	12.30
13.00 mL	2.03	**15.00 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
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

Stoichiometric (Equivalence) Point pH = 7

Example 3: Use the following observation table to determine the concentration of sulfuric acid.

10.0 mL of H_2SO_4 $_{(aq)}$ titrated by 0.0350 mol/L of KOH $_{(aq)}$				
	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 KOH 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.

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

$$2 \text{ KOH}_{(aq)} \quad + \quad H_2SO_{4\,(aq)} \quad \rightarrow \quad 2 \text{ HOH}_{(l)} \quad + \quad K_2SO_{4\,(aq)}$$
$$21.50 \text{ mL} \qquad\qquad 10.0 \text{ mL}$$
$$0.0350 \text{ mol/L} \qquad\quad ? \text{ mol/L}$$

$$[H_2SO_4] = 21.50 \cancel{\text{ mL KOH}} \times \frac{0.0350 \cancel{\text{ mol KOH}}}{1 \cancel{\text{ L KOH}}} \times \frac{1 \text{ mol } H_2SO_4}{2 \cancel{\text{ mol KOH}}} \times \frac{1}{10 \cancel{\text{ mL } H_2SO_4}} = 0.037625 \text{ mol/L}$$

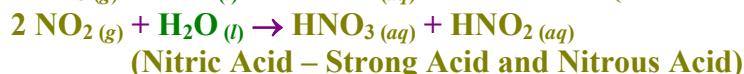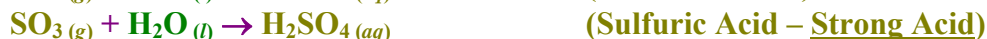$$\boxed{[H_2SO_4] = 0.0376 \text{ mol/L}}$$

Rainwater is Acidic and Ocean Water is Basic

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 CO_2 in the atmosphere.

$$CO_{2\ (g)} + H_2O_{\ (l)} \rightarrow H_2CO_{3\ (aq)} \qquad \textbf{(Carbonic Acid)}$$

- however, the burning of fossil fuel by vehicles and industrial activities, and also emits other oxides such as $NO_{2\ (g)}$, $SO_{2\ (g)}$ and $SO_{3\ (aq)}$. 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.
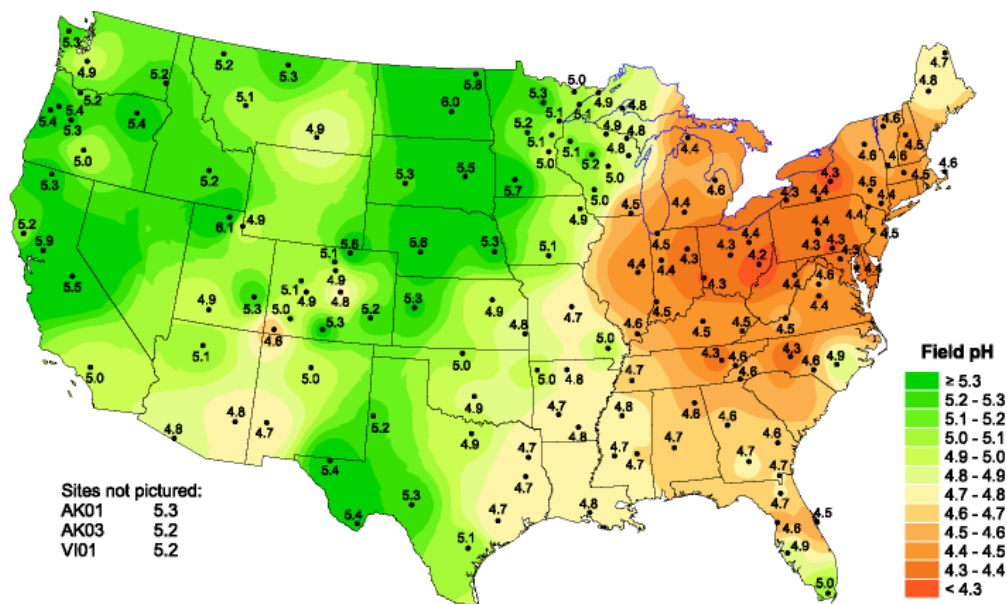
$$SO_{2\ (g)} + H_2O_{\ (l)} \rightarrow H_2SO_{3\ (aq)} \qquad \textbf{(Sulfurous Acid)}$$
$$SO_{3\ (g)} + H_2O_{\ (l)} \rightarrow H_2SO_{4\ (aq)} \qquad \textbf{(Sulfuric Acid – \underline{Strong Acid})}$$
$$2\ NO_{2\ (g)} + H_2O_{\ (l)} \rightarrow HNO_{3\ (aq)} + HNO_{2\ (aq)}$$
$$\textbf{(Nitric Acid – \underline{Strong Acid} and Nitrous Acid)}$$

- depending on the local rock formation, some areas, such as midwestern U.S., consist of a lot of limestone, calcium carbonate ($CaCO_3$), left by ocean water when the are was submerged millions of years ago. The limestone neutralizes acid rain, and one of the resulting product, carbon dioxide is released back into the atmosphere. However, other areas, like the northeastern U.S., consist of inert material such as granite, are affected more by acid rain.

$$2\ H_3O^+_{\ (aq)} + CaCO_{3\ (s)} \rightarrow 3\ H_2O_{\ (l)} + CO_{2\ (g)} + Ca^{2+}_{\ (aq)}$$

- in oceans, unlike rivers and lakes, contains a large amount of calcium carbonate. The carbon dioxide in the atmosphere is "deposited" in the ocean as carbonic acid ($CO_2 + H_2O$) is neutralized by the calcium carbonate. The resulting product of calcium hydrogen carbonate, a deposit on the ocean floor, makes the ocean water naturally basic (pH = 8.2). Hence, oceans are referred to as **"CO_2 sink"**.

$$CO_{2\ (g)} + H_2O_{\ (l)} \rightarrow H_2CO_{3\ (aq)} \qquad H_2CO_{3\ (aq)} + CaCO_{3\ (s)} \rightarrow Ca(HCO_3)_{2\ (s)}$$

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

Sites not pictured:
AK01   5.3
AK03   5.2
VI01   5.2

National Atmospheric Deposition Program/National Trends Network
http://nadp.sws.uiuc.edu

Field pH

≥ 5.3
5.2 - 5.3
5.1 - 5.2
5.0 - 5.1
4.9 - 5.0
4.8 - 4.9
4.7 - 4.8
4.6 - 4.7
4.5 - 4.6
4.4 - 4.5
4.3 - 4.4
< 4.3

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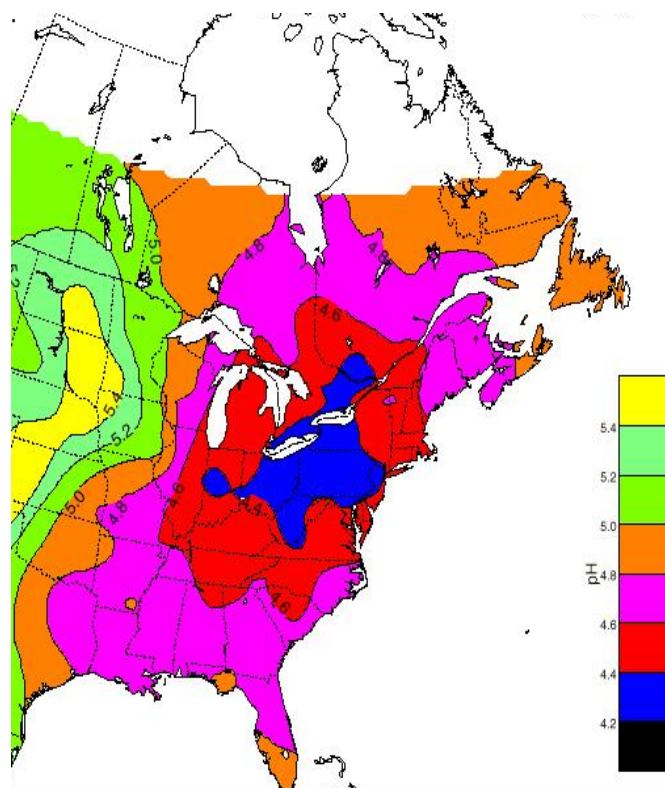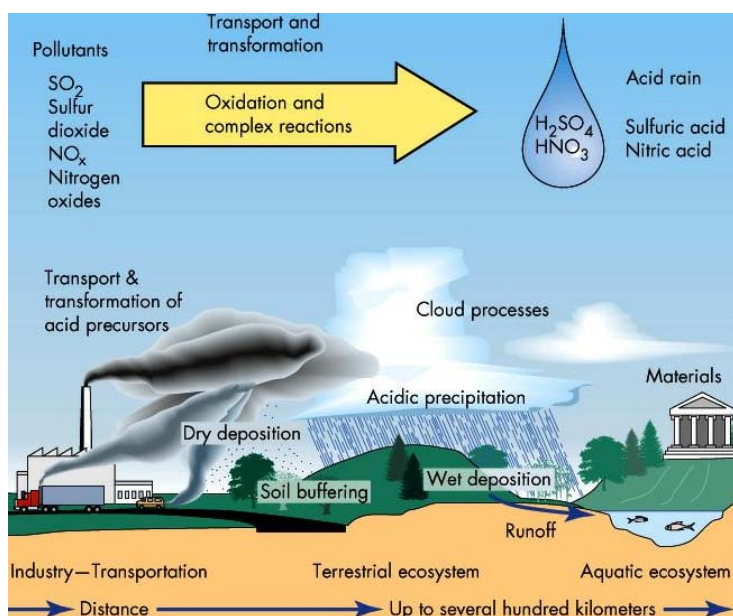
(Above) The Formation of Acid Rain.
(Right) Measurements of pH in Eastern Canada and
U.S in 1996. The lowest pHs are in the Great Lakes
Region where most of the heavy industries are located.

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 (SO_2 + H_2O) – Combination of *Sm*oke from Coal Burning and F*og* . 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, $CaCO_3$) 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 statutes.

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** (a process known as ***liming***) to acid lakes and rivers can raise the pH and reintroduce marine life that has been affected by acid rain.

> **Assignment**
> **15.3 pg. 556 #1 to 4 (Practice); pg. 556 #1 to 12**
> **Chapter 15 Review pg. 566−570 #16 to 18, 20 to 22, 24 to 26, 29 to 32, 41 to 75 (odd), 76 to 79, 81, and 97a**