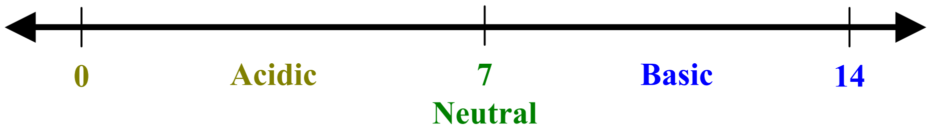
Chapter 10: Acids and Bases

10.1: Acids Donate Protons, Bases Accept Them

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).
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.
$pH < 7$	$pH > 7$

pH Scale



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

$$\boxed{H^+ + \text{Anion} \rightarrow \text{Acid}}$$

Example: $HCl_{(g)} \xrightarrow{\hspace{2cm}} HCl_{(aq)}$ $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$ ($HCl_{(aq)}$ is an Acid.)

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

Example: $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$ ($NaOH_{(aq)}$ is a Base.)

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 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 2: Provide chemical formulas for the following acids.

a. hydrocyanic acid

hydro*cyan*ic acid

hydrogen cyan*ide* ⇒ H^+ & CN^-

$HCN_{(aq)}$

b. oxalic acid

oxal*ic acid* ⇒ **hydrogen oxal*ate***

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

$H_2OOCCOO_{(aq)}$

c. sulfurous acid

sulfur*ous acid*

⇒ **hydrogen sulf*ite***

H^+ & SO_3^{2-}

$H_2SO_{3\,(aq)}$

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 3: 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 4: Provide chemical formulas for the following acids.

a. strontium hydroxide

Sr^{2+} and OH^- ⇒ $Sr(OH)_2$

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

b. rubidium hydroxide

Rb^+ and OH^- ⇒ $RbOH$

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

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.

$HBr_{(aq)} + H_2O_{(l)} \rightarrow \mathbf{H_3O^+_{(aq)}} + Br^-_{(aq)}$

(HBr is a Brønsted-Lowry Acid – donated a proton)

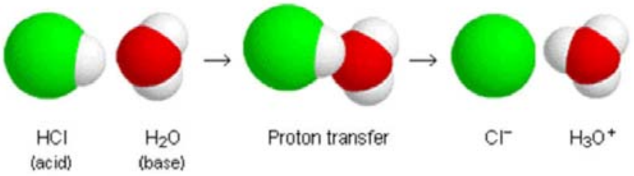
(H_2O is a Brønsted-Lowry Base – accepted a proton.)

$NH_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+_{(aq)} + \mathbf{OH^-_{(aq)}}$

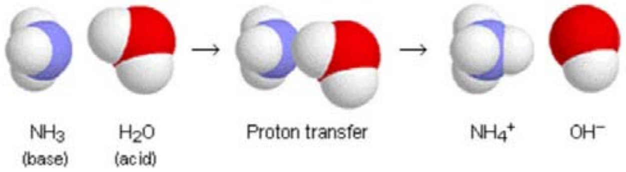
(H_2O is a Brønsted-Lowry Acid – donated a proton)

(NH_3 is a Brønsted-Lowry Base – accepted a proton.)

The acid HCl reacting with water. $HCl + H_2O \longrightarrow Cl^- + H_3O^+$



The base NH3 reacting with water. $NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$



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: $\mathbf{HA_{(aq)} + MOH_{(aq)} \rightarrow HOH_{(l)} + \boxed{MA_{(aq)}}}$

$\mathbf{H_2O_{(l)}}$

Examples:

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

Example 5: Write a balanced chemical equation for the following neutralization reactions.

a. hydrocyanic acid with lithium hydroxide

$\mathbf{HCN_{(aq)} + LiOH_{(aq)} \rightarrow HOH_{(l)} + LiCN_{(aq)}}$

$H^+\ CN^- \quad Li^+\ OH^-$

b. sulfuric acid with cesium hydroxide

$\mathbf{H_2SO_{4\,(aq)} + \underline{2}\ CsOH_{(aq)} \rightarrow \underline{2}\ HOH_{(l)} + Cs_2SO_{4\,(aq)}}$

$H^+\ SO_4^{2-} \quad Cs^+\ OH^-$

c. phosphoric acid with strontium hydroxide

$\mathbf{\underline{2}\ H_3PO_{4\,(aq)} + \underline{3}\ Sr(OH)_{2\,(aq)} \rightarrow \underline{6}\ HOH_{(l)} + Sr_3(PO_4)_{2\,(s)}}$

$H^+\ PO_4^{3-} \quad Sr^{2+}\ OH^-$

Assignment
10.1 pg. 332 #1 to 5

10.2: Some Acids and Bases are Stronger Than Others

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

- after dissociation is completed, the final acid concentration, $[HA]_f \approx 0$; $[H_3O^+]_{final} = [A^-]_{final} = [HA]_{original}$. **([ ] means concentration)**
- in the acid and base table (see next page), they are the ***ones ABOVE the grey line***.
- ***Note:*** **Strong Acids DO NOT MEAN that they are VERY CORROSIVE. It is the $[H_3O^+]$ that defines acidity.**

$$\underset{\text{(Strong Acid)}}{HA_{(aq)}} + \underset{\text{(Base)}}{H_2O_{(l)}} \longrightarrow H_3O^+_{(aq)} + A^-_{(aq)}$$

	$[HA]$		$[H_3O^+]$	$[A^-]$
Original	x		0	0
Change	$-x$		$+x$	$+x$
Final	0		x	x

Examples: Strong Acids: $HClO_{4\,(aq)}$, $HI_{(aq)}$, $HBr_{(aq)}$, $HCl_{(aq)}$, $H_2SO_{4\,(aq)}$ and $HNO_{3\,(aq)}$

Example 1: Write the dissociation reaction of 0.250 mol/L of $HBr_{(aq)}$ and determine its $[H_3O^+]$.

$HBr_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Br^-_{(aq)}$

(Strong Acid means $[HBr]_0 = [H_3O^+]_f$)

$[HBr]_0 = 0.250$ mol/L $\boxed{[H_3O^+]_f = 0.250 \text{ mol/L}}$

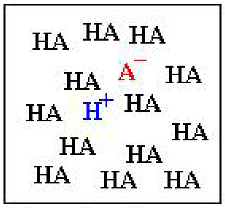
Weak Acids: - acids that dissociate LESS than 100% in water.

- when dissociation is completed, the hydronium concentration is much less than the original acid concentration, $[HA]_{final} > [H_3O^+]_{final}$ or $[HA]_{original} \approx [HA]_{final}$.
- in the acid and base table (see next page), they are the ***ones BELOW the grey line***.
- ***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.**

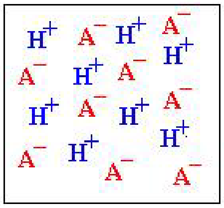
$$\underset{\text{(Weak Acid)}}{HA_{(aq)}} + \underset{\text{(Base)}}{H_2O_{(l)}} \overset{\longleftarrow}{\rightleftharpoons} H_3O^+_{(aq)} + A^-_{(aq)}$$

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

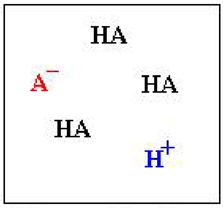
Examples: Some Weak Acids: $HOOCCOOH_{(aq)}$, $H_2SO_{3\,(aq)}$, $HSO_4^-{}_{(aq)}$, $H_3PO_{4\,(aq)}$, $HNO_{2\,(aq)}$, $HF_{(aq)}$,



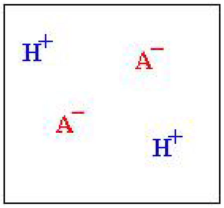
Concentrated weak acid - a lot present, but little dissociation of acid



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

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

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

Strong Acids

Very Weak Bases

Acid Name	Acid Formula	Base Formula
perchloric acid	$HClO_{4\,(aq)}$	$ClO_4^{-}{}_{(aq)}$
hydroiodic acid	$HI_{(aq)}$	$I^{-}{}_{(aq)}$
hydrobromic acid	$HBr_{(aq)}$	$Br^{-}{}_{(aq)}$
hydrochloric acid	$HCl_{(aq)}$	$Cl^{-}{}_{(aq)}$
sulfuric acid	$H_2SO_{4\,(aq)}$	$HSO_4^{-}{}_{(aq)}$
nitric acid	$HNO_{3\,(aq)}$	$NO_3^{-}{}_{(aq)}$
hydronium ion	$H_3O^{+}{}_{(aq)}$	$H_2O_{(l)}$
oxalic acid	$HOOCCOOH_{(aq)}$	$HOOCCOO^{-}{}_{(aq)}$
sulfurous acid ($SO_2 + H_2O$)	$H_2SO_{3\,(aq)}$	HSO_3^{-}
chlorous acid	$HClO_{2\,(aq)}$	$ClO_2^{-}{}_{(aq)}$
hydrogen sulfate ion	$HSO_4^{-}{}_{(aq)}$	SO_4^{2-}
phosphoric acid	$H_3PO_{4\,(aq)}$	$H_2PO_4^{-}{}_{(aq)}$
nitrous acid	$HNO_{2\,(aq)}$	$NO_2^{-}{}_{(aq)}$
arsenic acid	$H_3AsO_{4\,(aq)}$	$H_2AsO_4^{-}{}_{(aq)}$
monochloracetic acid	$HC_2H_2ClO_2$	$C_2H_2ClO_2^{-}{}_{(aq)}$
citric acid	$H_3C_6H_5O_{7\,(aq)}$	$H_2C_6H_5O_7^{-}{}_{(aq)}$
hydrofluoric acid	$HF_{(aq)}$	$F^{-}{}_{(aq)}$
methanoic (formic) acid	$HCOOH_{(aq)}$	$HCOO^{-}{}_{(aq)}$
hydrogen oxalate ion	$HOOCCOO^{-}{}_{(aq)}$	$OOCCOO^{2-}{}_{(aq)}$
lactic acid	$HC_3H_5O_{3\,(aq)}$	$C_3H_5O_3^{-}{}_{(aq)}$
ascorbic acid (vitamin C)	$H_2C_6H_6O_{6\,(aq)}$	$HC_6H_6O_6^{-}{}_{(aq)}$
benzoic acid	$C_6H_5COOH_{(aq)}$	$C_6H_5COO^{-}{}_{(aq)}$
ethanoic (acetic) acid	$CH_3COOH_{(aq)}$	$CH_3COO^{-}{}_{(aq)}$
dihydrogen citrate ion	$H_2C_6H_5O_7^{-}{}_{(aq)}$	$HC_6H_5O_7^{2-}{}_{(aq)}$
hydrated aluminum (III) ion	$[Al(H_2O)_6]^{3+}$	$[AlOH(H_2O)_5]^{2+}$
propanoic acid	$C_2H_5COOH_{(aq)}$	$C_2H_5COO^{-}{}_{(aq)}$
carbonic acid ($CO_2 + H_2O$)	$H_2CO_{3\,(aq)}$	$HCO_3^{-}{}_{(aq)}$
hydrogen citrate ion	$HC_6H_5O_7^{2-}{}_{(aq)}$	$C_6H_5O_7^{3-}{}_{(aq)}$
hydrosulfuric acid	$H_2S_{(aq)}$	$HS^{-}{}_{(aq)}$
dihydrogen arsenate ion	$H_2AsO_4^{-}{}_{(aq)}$	$HAsO_4^{2-}{}_{(aq)}$
hydrogen sulfite ion	$HSO_3^{-}{}_{(aq)}$	$SO_3^{2-}{}_{(aq)}$
dihydrogen phosphate ion	$H_2PO_4^{-}{}_{(aq)}$	$HPO_4^{2-}{}_{(aq)}$
hypochlorous acid	$HOCl_{(aq)}$	$OCl^{-}{}_{(aq)}$
hypobromous acid	$HOBr_{(aq)}$	$OBr^{-}{}_{(aq)}$
hydrocyanic acid	$HCN_{(aq)}$	$CN^{-}{}_{(aq)}$
hydrogen arsenate ion	$HAsO_4^{2-}{}_{(aq)}$	$AsO_4^{3-}{}_{(aq)}$
boric acid	$H_3BO_{3\,(aq)}$	$H_2BO_3^{-}{}_{(aq)}$
ammonium ion	$NH_4^{+}{}_{(aq)}$	$NH_{3\,(aq)}$
phenol	$C_6H_5OH_{(aq)}$	$C_6H_5O^{-}{}_{(aq)}$
hydrogen carbonate ion	$HCO_3^{-}{}_{(aq)}$	$CO_3^{2-}{}_{(aq)}$
hypoiodous acid	$HOI_{(aq)}$	$OI^{-}{}_{(aq)}$
hydrogen ascorbate ion	$HC_6H_6O_6^{-}{}_{(aq)}$	$C_6H_6O_6^{2-}{}_{(aq)}$
hydrogen phosphate ion	$HPO_4^{2-}{}_{(aq)}$	$PO_4^{3-}{}_{(aq)}$
hydrogen sulfide ion	$HS^{-}{}_{(aq)}$	$S^{2-}{}_{(aq)}$
water (55.49 mol/L)	$H_2O_{(l)}$	$OH^{-}{}_{(aq)}$

Increasing Acid Strength ↑

Increasing Base Strength ↓

Strong Bases: - bases that **dissociate completely (100%) in water**.
- all **alkali bases (Group IA cations with OH^-)** and some **alkaline bases (Group IIA cations with OH^-)** are considered as strong bases because they are ionic compound that dissociates completely.

Examples: Strong Ionic Bases:

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

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)} \qquad KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$$

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

$$Ba(OH)_{2\,(aq)} \rightarrow Ba^{2+}_{(aq)} + 2\,OH^-_{(aq)} \qquad Sr(OH)_{2\,(aq)} \rightarrow Sr^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

Example 2: Write the dissociation reaction of 0.420 mol/L of $Sr(OH)_{2\,(aq)}$ and determine its $[OH^-]$.

$$Sr(OH)_{2\,(aq)} \rightarrow Sr^{2+}_{(aq)} + \underline{2}\,OH^-_{(aq)}$$

There are 2 OH^- dissociated for each $Sr(OH)_2$ dissolves.

0.420 mol/L 2 × 0.420 mol/L $\boxed{[OH^-] = 0.840 \text{ mol/L}}$

Weak Bases: - bases that **dissociate LESS than 100% in water**.
- after dissociation is completed, the hydroxide concentration is much less than the original base concentration, $[OH]_{final} < [B]_{final}$ or $[B]_{original} \approx [B]_{final}$.
- in the acid and base table (see last page), they are the ***ones ABOVE*** **OH^-**.

(Weak Base) **(Acid)** ←

$$B_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HB^+_{(aq)}$$

	$[B]$		$[OH^-]$	$[HB^+]$
Initial **Change**	x $-y$ (where $y << x$)		0 $+y$	0 $+y$
Final	$(x - y) \approx x$		y	y

Examples: Some Weak Bases: $HOOCCOO^-_{(aq)}$, $HSO_3^-{}_{(aq)}$, $HSO_4^-{}_{(aq)}$, $H_2PO_4^-{}_{(aq)}$, $NO_2^-{}_{(aq)}$, $HCO_3^-{}_{(aq)}$, $H_2C_6H_5O_7^-{}_{(aq)}$, $F^-_{(aq)}$, $HCOO^-_{(aq)}$, $C_6H_7O_6^-{}_{(aq)}$, $C_6H_5COO^-_{(aq)}$, $CO_3^{2-}{}_{(aq)}$ $CH_3COO^-_{(aq)}$, $HS^-_{(aq)}$, $OCl^-_{(aq)}$, $CN^-_{(aq)}$, $NH_{3\,(aq)}$, and $NO_3^-{}_{(aq)}$

Example 3: 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:

$$HClO_{4\,(aq)} >> H_2SO_{3\,(aq)} > HNO_{2\,(aq)} > HF_{(aq)}$$

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

$NaCN_{(aq)}$, $Na_2SO_{4\,(aq)}$, $NaCl_{(aq)}$, $NaCH_3COO_{(aq)}$, $NaOH_{(aq)}$

All compounds above dissociates where the anion is the base.

$$NaCN_{(aq)} \rightarrow Na^+_{(aq)} + CN^-_{(aq)} \qquad Na_2SO_{4\,(aq)} \rightarrow \underline{2}\,Na^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

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

$$OH^-_{(aq)} >> CN^-_{(aq)} > CH_3COO^-_{(aq)} > SO_4^{2-}{}_{(aq)} >> Cl^-_{(aq)}$$

Assignment
10.2 pg. 332 #6 to 10

10.3: Solutions Can Be Acidic, Basic, or Neutral

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.

$HA_{(aq)}$	+	**$H_2O_{(l)}$**	⇌	**$H_3O^+_{(aq)}$**	+	**$A^-_{(aq)}$**
(Acid)		**(Base)**		**(Conjugate Acid)**		**(Conjugate Base)**
$B_{(aq)}$	+	**$H_2O_{(l)}$**	⇌	**$HB^+_{(aq)}$**	+	**$OH^-_{(aq)}$**
(Base)		**(Acid)**		**(Conjugate Acid)**		**(Conjugate Base)**

Examples:

$HBr_{(aq)}$	+	$H_2O_{(l)}$	→	$H_3O^+_{(aq)}$	+	$Br^-_{(aq)}$
(Acid)		(Base)		(Conjugate Acid)		(Conjugate Base)
$NH_{3\,(aq)}$	+	$H_2O_{(l)}$	⇌	$NH_4^+{}_{(aq)}$	+	$OH^-_{(aq)}$
(Base)		(Acid)		(Conjugate Acid)		(Conjugate Base)

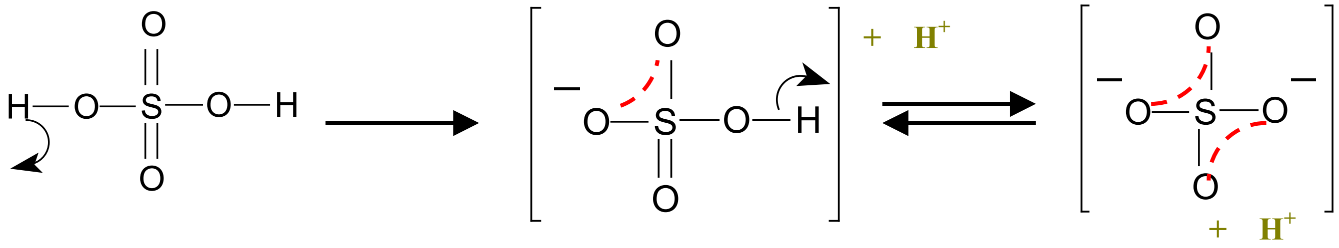
Monoprotic Acids: - acids that can donate a maximum of one proton.

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).

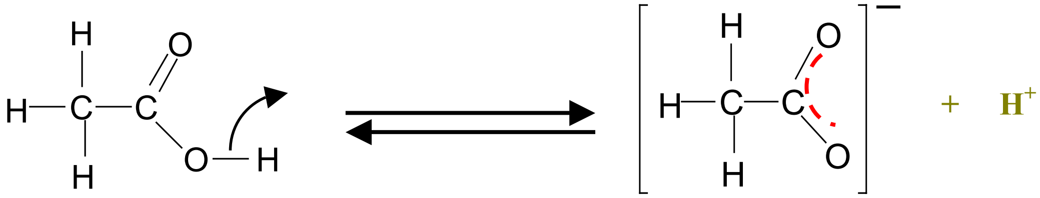
$H_2A_{(aq)}$	+	**$H_2O_{(l)}$**	⇌	**$H_3O^+_{(aq)}$**	+	**$HA^-_{(aq)}$**
(Original Diprotic Acid)		**(Base)**		**(Conjugate Acid)**		**(Conjugate Base)**
$HA^-_{(aq)}$	+	**$H_2O_{(l)}$**	⇌	**$H_3O^+_{(aq)}$**	+	**$A^{2-}_{(aq)}$**
(Acid)		**(Base)**		**(Conjugate Acid)**		**(Conjugate Base)**

(Conjugate Base of Original Diprotic Acid)

Example: Sulfuric Acid ($H_2SO_{4\,(aq)}$), successively donate protons to finally reach $SO_4^{2-}{}_{(aq)}$



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



Triprotic Acids: - acids that can donate a maximum of three protons in stepwise dissociation.
- the conjugate base of the original acid can itself be an acid.

$$H_3A_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + H_2A^-_{(aq)}$$

(Original Triprotic Acid) (Base) (Conjugate Acid) (Conjugate Base)

$$H_2A^-_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HA^{2-}_{(aq)}$$

(Acid) (Base) (Conjugate Acid) (Conjugate Base)

(Conjugate Base of Original Triprotic Acid)

$$HA^{2-}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^{3-}_{(aq)}$$

(Acid) (Base) (Conjugate Acid) (Conjugate Base)

(Conjugate Base after 2nd Proton Donation)

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: $HOOCCOO^-_{(aq)}$, $HSO_4^-{}_{(aq)}$, $HSO_3^-{}_{(aq)}$, $HCO_3^-{}_{(aq)}$, $HS^-_{(aq)}$, $HC_6H_6O_6^-{}_{(aq)}$, $H_2PO_4^-{}_{(aq)}$, $HPO_4^-{}_{(aq)}$, and $\mathbf{H_2O}_{(l)}$

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

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)**, $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/L

Autoionization of Water

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

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

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

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

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

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

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0\times10^{-14}}{1.0\times10^{-4}}$$

$$\boxed{[H_3O^+] = 1.0 \times 10^{-10} \text{ mol/L}}$$

b. $[H_3O^+] = 5.0 \times 10^{-5}$ M

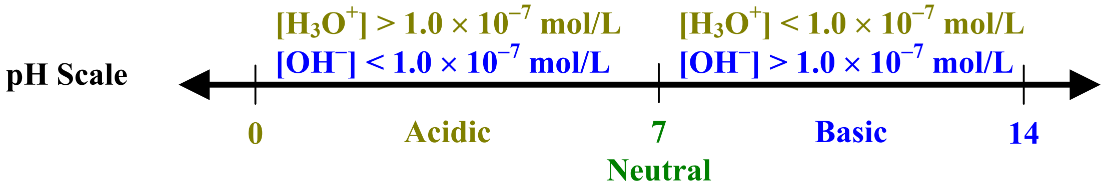
$$K_w = [H_3O^+][OH^-]$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{5.0\times10^{-5}}$$

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

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 $[OH^-]$**).
- 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.**



Acidic Solution: - solutions that has a **pH < 7**, where **$[H_3O^+] > 1.0 \times 10^{-7}$ M** and **$[OH^-] < 1.0 \times 10^{-7}$ M**.

Basic Solution: - solutions that has a **pH > 7**, where **$[H_3O^+] < 1.0 \times 10^{-7}$ M** and **$[OH^-] > 1.0 \times 10^{-7}$ M**.

- also known as **Alkaline Solution** (basic solutions that were from alkali metal hydroxides (LiOH, NaOH, KOH, RbOH, and CsOH) or alkaline metal hydroxides ($Ca(OH)_2$, $Ba(OH)_2$, and $Sr(OH)_2$)

Neutral Solution: - solution that has a **pH = 7**, where **$[H_3O^+] = 1.0 \times 10^{-7}$ M** and **$[OH^-] = 1.0 \times 10^{-7}$ M.**

- compounds that do **NOT** have any weak acid and base as cations or anions.
- however, very, very weak acids or bases as cations or anions are acceptable. This is because extremely weak acids or bases (such as Cl^-, NO_3^-, ClO_4^-, Br^- and I^-) cannot donate or accept any protons.

Examples: NaCl, KBr, $Ba(NO_3)_2$, LiI, and $Sr(ClO_4)_2$

pH of Some Common Substances

Substance	pH
1 M of HCl	0.00
Battery Acid	0.00
Stomach Acid	2.00
Lemon Juice	2.50
Vinegar	2.50
Soft drink	2.80
Beer	3.20

Substance	pH
Tomato	3.90
Coffee	4.30
Milk	6.30
Saliva	6.50
Clean Rain Water	6.70
Pure Water	7.00
Blood	7.50

Substance	pH
Sea Water	8.00
Baking Soda ($NaHCO_3$)	9.00
Soap	9.68
Ammonia as Cleaner	10.50
Hair Remover	12.00
Oven Cleaner	13.00
1 M of NaOH	14.00

pH Scale

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

Example 2: Calculate the pH for the following solutions.

a. $[H^+] = 0.0100$ mol/L

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

b. $[H_3O^+] = 5.00 \times 10^{-6}$ M

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

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

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

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

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0\times10^{-14}}{1.50}$$

$[H_3O^+] = 6.67 \times 10^{-15}$ mol/L

$pH = -\log[H_3O^+]$
$pH = -\log(6.67 \times 10^{-15})$
$\boxed{pH = 14.18}$

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

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

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

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0\times10^{-14}}{2.00\times10^{-3}}$$

$[H_3O^+] = 5.00 \times 10^{-12}$ mol/L

$pH = -\log[H_3O^+]$
$pH = -\log(5 \times 10^{-12})$
$\boxed{pH = 11.30}$

Example 3: Calculate the $[H_3O^+]$ and the $[OH^-]$ for the following solutions.

a. pH = 4.00

$pH = -\log[H_3O^+]$
$[H_3O^+] = 10^{-pH} = 10^{-4.00}$
$\boxed{[H_3O^+] = 1.00 \times 10^{-4}\text{ M}}$

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

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{1.0\times10^{-4}}$$

$\boxed{[OH^-] = 1.0 \times 10^{-10}\text{ M}}$

b. pH = 9.67

$pH = -\log[H_3O^+]$
$[H_3O^+] = 10^{-pH} = 10^{-9.67}$
$\boxed{[H_3O^+] = 2.14 \times 10^{-10}\text{ M}}$

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

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{2.14\times10^{-10}}$$

$\boxed{[OH^-] = 4.67 \times 10^{-5}\text{ M}}$

c. pH = 12.83

$pH = -\log[H_3O^+]$
$[H_3O^+] = 10^{-pH} = 10^{-12.83}$
$\boxed{[H_3O^+] = 1.48 \times 10^{-13}\text{ M}}$

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

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0\times10^{-14}}{1.48\times10^{-13}}$$

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

Acid-Base Indicators:
- chemicals that change colours at a specific pH range.
- they are themselves weak organic acids. Since they are usually very big structurally, we use abbreviations to describe them in chemical equations.
- the acidic form of the indictor has a different colour than its basic form.

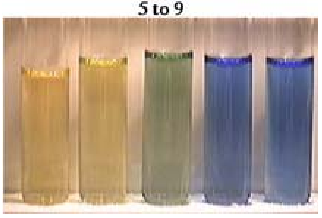
Example: Bromothymol Blue ($HBb_{(aq)}$ /$Bb^-_{(aq)}$) changes colours from yellow to blue at a pH range of 6.0 to 7.6. At its **mid-range (6.8 – average of 6.0 and 7.6)**, the colour would be **green (yellow + blue)**.

$$\underset{\textbf{(Yellow)}}{HBb_{(aq)}} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + \underset{\textbf{(Blue)}}{Bb^-_{(aq)}}$$

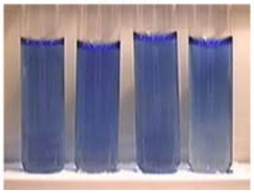
At pH ≤ 6, bromothymol blue is yellow. At pH ≥ 8, bromothymol blue is blue.

1 to 4

5 to 9



10 to 13



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	$H_2Cr_{(aq)}$ / $HCr^{-}_{(aq)}$ $HCr_{(aq)}$ / $Cr^{2-}_{(aq)}$	0.0 – 1.0 7.0 – 8.8	red to yellow yellow to red
thymol blue	$H_2Tb_{(aq)}$ / $HTb^{-}_{(aq)}$ $HTb^{-}_{(aq)}$ / $Tb^{2-}_{(aq)}$	1.2 – 2.8 8.0 – 9.6	red to yellow 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

Example 4: Using the indicator table above, complete the following table for a solution has a pH of 4.5.

Indicators	Observed Colours
methyl violet	**blue**
methyl orange	**yellow**
bromocresol green	**green**
methyl red	**red**
phenolphthalein	**colourless**

Assignment

10.3 pg. 332 #11 to 17 and Worksheet: Acids and Bases
Hands-On Activity: Rainbow Cabbage (pg. 323)

Worksheet: Acids and Bases

1. Name or give the formulas of the following acids and classified whether they are monoprotic, diprotic or triprotic.

a. $H_2SO_3\,_{(aq)}$
b. $HNO_3\,_{(aq)}$
c. $H_3BO_3\,_{(aq)}$
d. $HC_6H_5COO\,_{(aq)}$
e. $H_2Se\,_{(aq)}$
f. hydrosulfuric acid
g. hypochlorous acid
h. sulfuric acid
i. hydroarsenic acid
j. carbonic acid

2. Using the *Relative Strengths of Acid and Base Table* on pg. 123, write the Brønsted-Lowry dissociation equations for the following chemicals. State whether they are acids, bases or amphoteric species.

a. $HNO_2\,_{(aq)}$
b. $SO_3^{2-}\,_{(aq)}$
c. $HClO_2\,_{(aq)}$
d. $HPO_4^{2-}\,_{(aq)}$
e. $CO_3^{2-}\,_{(aq)}$
f. $HI\,_{(aq)}$
g. $H_2S\,_{(aq)}$
h. $F^{-}\,_{(aq)}$
i. $HSO_4^{-}\,_{(aq)}$
j. $NH_4^{+}\,_{(aq)}$

3. Write a balanced equation for the following acid and base neutralization reactions.

a. nitric acid and sodium hydroxide
b. hydrosulfuric acid and potassium hydroxide
c. benzoic acid and barium hydroxide
d. carbonic acid and strontium hydroxide
e. phosphoric acid and lithium hydroxide
f. boric acid and calcium hydroxide

4. Given the initial concentrations of the following strong acids and bases, determine the $[H^+]$ and $[OH^-]$ as well as the pH.

a. 0.200 mol/L of $HI\,_{(aq)}$
b. 0.012 M of $KOH\,_{(aq)}$
c. 0.0075 mol/L of $HClO_4\,_{(aq)}$
d. 0.075 M $LiOH\,_{(aq)}$
e. 4.5×10^{-6} mol/L of $HNO_3\,_{(aq)}$
f. 8.25×10^{-4} M $Sr(OH)_2\,_{(aq)}$

5. Determine the $[H_3O^+]$ and $[OH^-]$ with the following pH.

a. pH = 3.00
b. pH = 10.00
c. pH = 4.15
d. pH = 11.86
e. pH = 6.71
f. pH = 13.4

Use the following information to answer the next question.

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	?

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

7. A base solution is slowly added to an acidic solution in such a way that the $[H_3O^+]$ is changing from 1×10^{-3} M to 1×10^{-4} M. Calculate the pH range of this change. List several indicators and their colours that can be used to signal when the mixing should stop.

Answers

1a. sulfurous acid; diprotic
b. nitric acid; monoprotic
c. boric acid; triprotic
d. benzoic acid; monoprotic
e. hydroselenic acid; diprotic
f. $H_2S_{(aq)}$; diprotic
g. $HClO_{(aq)}$; monoprotic
h. $H_2SO_{4(aq)}$; diprotic
i. $H_3As_{(aq)}$; triprotic
j. $H_2CO_{3(aq)}$; diprotic

2a. acid; $HNO_{2(aq)} + H_2O_{(l)} \rightleftharpoons {H_3O}^{+}_{(aq)} + {NO_2}^{-}_{(aq)}$
b. base; ${SO_3}^{2-}_{(aq)} + H_2O_{(l)} \rightleftharpoons {HSO_3}^{-}_{(aq)} + OH^{-}_{(aq)}$
c. acid; $HClO_{2(aq)} + H_2O_{(l)} \rightleftharpoons {H_3O}^{+}_{(aq)} + {ClO_2}^{-}_{(aq)}$
d. amphoteric; (acid): ${HPO_4}^{2-}_{(aq)} + H_2O_{(l)} \rightleftharpoons {H_3O}^{+}_{(aq)} + {PO_4}^{3-}_{(aq)}$
(base): ${HPO_4}^{2-}_{(aq)} + H_2O_{(l)} \rightleftharpoons {H_2PO_4}^{-}_{(aq)} + OH^{-}_{(aq)}$
e. base; ${CO_3}^{2-}_{(aq)} + H_2O_{(l)} \rightleftharpoons {HCO_3}^{-}_{(aq)} + OH^{-}_{(aq)}$
f. acid; $HI_{(aq)} + H_2O_{(l)} \rightleftharpoons {H_3O}^{+}_{(aq)} + I^{-}_{(aq)}$
g. acid; $H_2S_{(aq)} + H_2O_{(l)} \rightleftharpoons {H_3O}^{+}_{(aq)} + S^{2-}_{(aq)}$
h. base; $F^{-}_{(aq)} + H_2O_{(l)} \rightleftharpoons HF_{(aq)} + OH^{-}_{(aq)}$
i. amphoteric; (acid): ${HSO_4}^{-}_{(aq)} + H_2O_{(l)} \rightleftharpoons {H_3O}^{+}_{(aq)} + {SO_4}^{2-}_{(aq)}$
(base): ${HSO_4}^{-}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_2SO_{4(aq)} + OH^{-}_{(aq)}$
j. acid; ${NH_4}^{+}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + {H_3O}^{+}_{(aq)}$

3a. $HNO_{3(aq)} + NaOH_{(aq)} \rightarrow HOH_{(l)} + NaNO_{3(aq)}$ **b.** $H_2S_{(aq)} + 2\ KOH_{(aq)} \rightarrow 2\ HOH_{(l)} + K_2S_{(aq)}$
c. $2\ HC_6H_5COO_{(aq)} + Ba(OH)_{2(aq)} \rightarrow 2\ HOH_{(l)} + Ba(C_6H_5COO)_{2(aq)}$
d. $H_2CO_{3(aq)} + Sr(OH)_{2(aq)} \rightarrow 2\ HOH_{(l)} + SrCO_{3(s)}$ **e.** $H_3PO_{4(aq)} + 3\ LiOH_{(aq)} \rightarrow 3\ HOH_{(l)} + Li_3PO_{4(aq)}$
f. $2\ H_3BO_{3(aq)} + 3\ Ca(OH)_{2(aq)} \rightarrow 6\ HOH_{(l)} + Ca_3(BO_3)_{2(s)}$

4a. $[H^+] = 0.200$ mol/L; $[OH^-] = 5.00 \times 10^{-14}$ mol/L; pH = 0.699
b. $[H^+] = 8.33 \times 10^{-13}$ M; $[OH^-] = 0.012$ M; pH = 12.08
c. $[H^+] = 0.0075$ mol/L; $[OH^-] = 1.33 \times 10^{-12}$ mol/L; pH = 2.12
d. $[H^+] = 1.33 \times 10^{-13}$ mol/L; $[OH^-] = 0.075$ M; pH = 12.88
e. $[H^+] = 4.5 \times 10^{-6}$ mol/L; $[OH^-] = 2.22 \times 10^{-9}$ mol/L; pH = 5.35
f. $[H^+] = 6.06 \times 10^{-12}$ M; $[OH^-] = 1.65 \times 10^{-3}$ M; pH = 11.22

5a. $[H_3O^+] = 1.00 \times 10^{-3}$ mol/L; $[OH^-] = 1.00 \times 10^{-11}$ mol/L
b. $[H_3O^+] = 1.00 \times 10^{-10}$ M; $[OH^-] = 1.00 \times 10^{-4}$ M
c. $[H_3O^+] = 7.08 \times 10^{-5}$ mol/L; $[OH^-] = 1.41 \times 10^{-10}$ mol/L
d. $[H_3O^+] = 1.38 \times 10^{-12}$ M; $[OH^-] = 7.24 \times 10^{-3}$ M
e. $[H_3O^+] = 1.95 \times 10^{-7}$ mol/L; $[OH^-] = 5.13 \times 10^{-8}$ mol/L
f. $[H_3O^+] = 3.98 \times 10^{-14}$ M; $[OH^-] = 0.251$ M

6. pH = 6.0 to 6.6; bromothymol blue will be yellow to green at that pH.

7. pH range between 3.00 to 4.00; possible indicators are methyl orange (starts out red, but stop at orange) and bromocresol green (starts out yellow, but stop at light green).

10.4: 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.

$$SO_{2\,(g)} + H_2O_{(l)} \rightarrow H_2SO_{3\,(aq)} \qquad \textbf{(Sulfurous Acid)}$$
$$SO_{3\,(g)} + H_2O_{(l)} \rightarrow H_2SO_{4\,(aq)} \qquad \textbf{(Sulfuric Acid – \underline{Strong Acid})}$$
$$2\,NO_{2\,(g)} + H_2O_{(l)} \rightarrow HNO_{3\,(aq)} + HNO_{2\,(aq)}$$

(Nitric Acid – Strong Acid and Nitrous Acid)

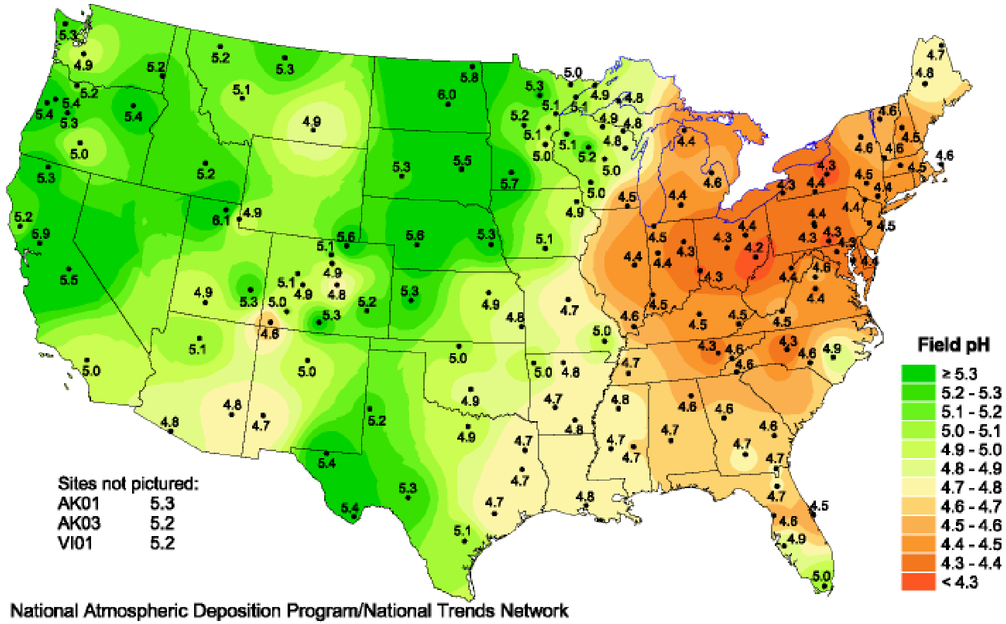
- depending on the local rock formation, some areas, such as midwestern U.S., consist of a lot of limestone, calcium carbonate ($CaCO_3$), left by ocean water when the are was submerged millions of years ago. The limestone neutralizes acid rain, and one of the resulting product, carbon dioxide is released 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**".

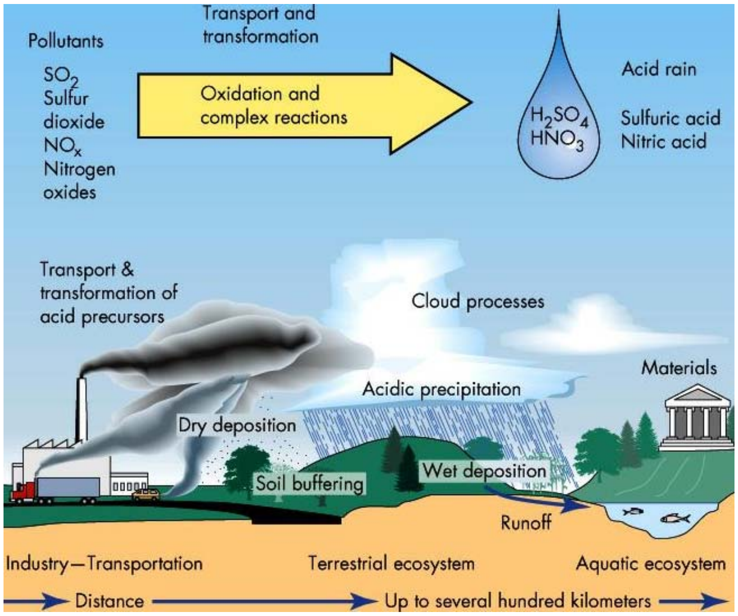
$$CO_{2\,(g)} + H_2O_{(l)} \rightarrow H_2CO_{3\,(aq)} \qquad H_2CO_{3\,(aq)} + CaCO_{3\,(s)} \rightarrow Ca(HCO_3)_{2\,(s)}$$

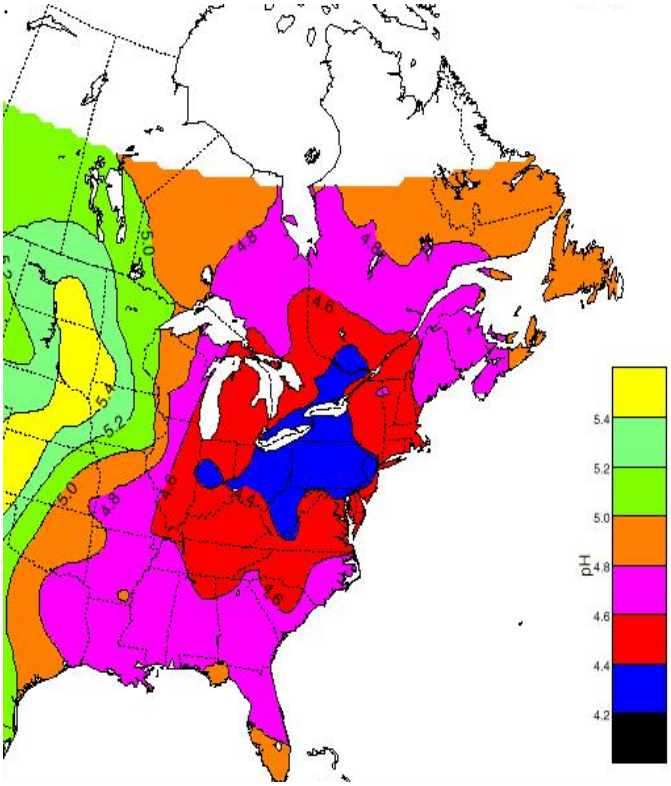
Hydrogen ion concentration as pH from measurements made at the field laboratories, 1999



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

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.

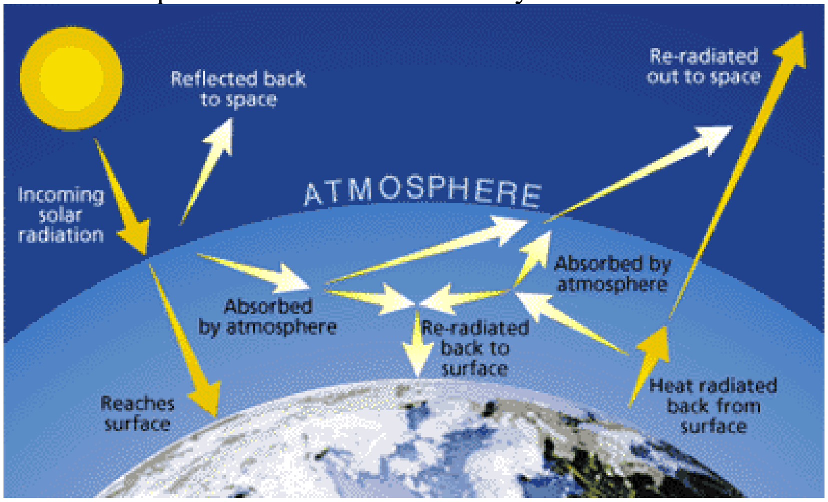
Ways to Reduce Acid Rain:

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.

Greenhouse Effect: - the emission of greenhouses gases that traps more of the sun's radiant (heat) energy in the atmosphere than it occurs naturally.



Greenhouses Gases: - man-made and naturally occur gases that contribute to the Greenhouse Effect.

1. **Carbon dioxide (CO_2)** is released to the atmosphere when solid waste, fossil fuels (oil, natural gas, and coal), and wood and wood products are burned.
2. **Methane (CH_4)** is emitted during the production and transport of coal, natural gas, and oil. Methane emissions also result from the decomposition of organic wastes in municipal solid waste landfills, and the raising of livestock.
3. **Nitrous oxide** (N_2O) is emitted during agricultural and industrial activities, as well as during combustion of solid waste and fossil fuels.
4. **Hydrofluorocarbons (HFCs), Perfluorocarbons (PFCs), and Sulfur Hexafluoride (SF_6)** are very powerful greenhouse gases that are not naturally occurring. They are generated by various industrial processes.

Each greenhouse gas differs in its ability to absorb heat in the atmosphere. HFCs and PFCs are the most heat-absorbent. Methane traps over 21 times more heat per molecule than carbon dioxide, and nitrous oxide absorbs 270 times more heat per molecule than carbon dioxide. Often, estimates of greenhouse gas emissions are presented in units of millions of metric tons of carbon equivalents (MMTCE), which weights each gas by its GWP value, or Global Warming Potential. (Information from US. EPA)

- Automobiles and Major Transportations account for 34% of CO_2 emissions globally (Power Plants contributes 33%; Major Industries and Home Heating contribute the remaining 33%).
- Presently 89% of Energy Productions involve the burning of Fossil Fuels (Coal, Petroleum, Natural Gas and Biomass).
- Heat and Electricity generated from combustion of fossil fuel is at most 30% efficient.

(Data from University of Michigan: http://www.umich.edu/~gs265/society/greenhouse.htm)

Deforestation: - another cause of an increased in CO_2 level in the atmosphere. As forests disappeared, there is a lack of plants to absorb carbon dioxide using photosynthesis.
- also causes mud and landslides, demineralization of the soil, lost animal habitats and extinction, destruction of entire ecosystems. Plants that may have important medicinal values can also be destroyed.

The Environmental Effect of Using Fossil Fuel: (Greenhouse Effect)

1. **Global Warming**: - the warming of global temperature due to an increase of greenhouse gases in the atmosphere.
2. **Rise of Water Level**: - low-lying islands and coastal area are endangered as polar icecaps melt due to the rise of temperature as a result of the greenhouse effect.
3. **Unpredicted and Erratic Climate**: - greenhouse effect is related to droughts and dry whether in many parts of the world.

Alternate Energy Sources without the Emission of Greenhouse Gas

1. **Solar Energy**: - the most efficient energy source where energy from the sun is converted directly to electricity through the use of photovoltaic cells (solar panels) or heat using high efficient insulated glass and an effective water heating system.
   - technology exists but expensive; requires many solar panels to generate adequate amount of electricity.
2. **Wind Power**: - the use of wind turbines to generate electricity.
   - very efficient and extremely economical, but location specific and not very reliable when there is no wind.
   - can disrupt migratory routes of birds (they get caught in the turbine), aesthetic problems for the landscapes.
3. **Geothermal Power**: - the use of underground steam to generate electricity.
   - very efficient somewhat reliable, but location specific.
   - geothermal power is widely use in Iceland where it is sitting on the Atlantic ridge and there are lots of hot springs.
4. **Tidal Power**: - the use of tidal current to generate electricity.
   - very efficient somewhat reliable, but location specific.
   - tidal power involves establishing electric turbines at a narrow mouth of a channel where normal tides can cause bigger current and quick rise in water levels. It is being used in the Bay of Fundy at Nova Scotia, Canada and Kvalsund at the Arctic tip of Norway.
   - tidal power can sometimes disrupt migratory routes of marine species.
5. **Hydroelectricity**: - the use of dam and reservoir to turn electric turbines as water falls from a higher level to the spillway (potential energy converted to kinetic energy to electricity).
   - very efficient and no emission of greenhouse gas.
   - location specific and very expensive to built. The reservoir flooding can destroy ecological habitats and force migrations of people from towns and villages (Aswan Dam in Egypt and the Three Gorges Dam in China displaced thousands of people and submerged ancient cities). The presence of the dam can disrupt aquatic migratory routes as well.
   - dams have a limited life span (the collection of silt and mud at the bottom of the dam has to be clear periodically to maintain the structural integrity of the dam). Dams can burst during earthquakes or poor maintenance. Flash flooding of towns along spillway is always a danger.
6. **Hydrogen Fuel**: - burning hydrogen to form water and generate heat and electricity.
   - very efficient and zero pollution.
   - hydrogen is very explosive and technologies are still needed for supplying and storing hydrogen safely in automobiles and homes.

Assignment

10.4 pg. 332–333 #18 to 23

10.5: Buffer Solution Resist Changes in pH

Buffer Solution: - a solution that **resists a change in pH** when a small amount of H_3O^+ or OH^- is added.
- consists of a pair of **weak acid (HA) / conjugate base common ion (A^-)** or a pair of **weak base/conjugate acid common ion**.

In a Buffer Solution: (HA $_{(aq)}$, $A^-_{(aq)}$, and $H_2O_{(l)}$)

a. **Small Amounts of H_3O^+ is Added:** $A^-_{(aq)} + H_3O^+_{(aq)} \rightarrow H_2O_{(l)} + HA_{(aq)}$

(H_3O^+ - Strong Acid reacts completely with A^-)

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

b. **Small Amounts of OH^- is Added:** $HA_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + A^-_{(aq)}$

(OH^- - Strong Base reacts completely with HA)

(More A^- – Weak Base: pH will only be raised SLIGHTLY!)

Example: Blood contains a $H_2CO_{3\,(aq)}$ / HCO_3^- buffer solution. The carbonic acid comes from when CO_2 produced during the process of cellular respiration enters the blood stream and reacts with water. Because H_2CO_3 acts as an acid and HCO_3^- behaves like 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: $H_2CO_{3\,(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + HCO^-_{3\,(aq)}$

When small amounts of strong acid is added: $HCO^-_{3\,(aq)} + H_3O^+_{(aq)} \rightarrow H_2O_{(l)} + H_2CO_{3(aq)}$

Holding one's Breath: - causes the pH of the body to decrease. The amount of $H_2CO_{3\,(aq)}$ increases because more CO_2 is retained in the body.

Hyperventilation: - when the breathing rate increased suddenly. The body quickly gets rid CO_2 causing the amount of H_2CO_3 to decrease. Hence, the pH of the body increases.
- whenever there is a period of lack of oxygen or an sudden decrease in pH (such as taking too much Asprin – a weak acid), the body responds by hyperventilating to try to rebalance the pH level.

Example 1: Write the chemical equations when small amounts of acids and bases are added separately to a buffered solution containing $CH_3COOH_{(aq)}$ and $NaCH_3COO_{(aq)}$ (dissociates to $Na^+_{(aq)}$ and $CH_3COO^-_{(aq)}$).

Buffer Solution: contains $CH_3COOH_{(aq)}$, $CH_3COO^-_{(aq)}$, and $H_2O_{(l)}$

When small amounts of strong acid is added: $CH_3COO^-_{(aq)} + H_3O^+_{(aq)} \rightarrow H_2O_{(l)} + CH_3COOH_{(aq)}$

When small amounts of strong base is added: $CH_3COOH_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + CH_3COO^-_{(aq)}$

Check out the Acid-Base Buffer Animation at
(http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/buffer12.swf)

Assignment

10.5 pg. 333 #24 to 28
Chapter 10 Review: pg. 332 #1 to 12 (Key Terms and Matching Definitions)
pg. 333–334 (Exercises) #1 to 28; pg. 334–335 (Problems) #1 to 5