ACIDS and BASES

1. a) Definition -Latin (acidus) = sour, vinegar and lemons.

   Arrhenius definition – Acid produces H^+

   $HX \rightarrow H^+_{(aq)} + X^-_{(aq)}$ $HBr \rightarrow H^+_{(aq)} + Br^-_{(aq)}$

   b) Nomenclature-

	Prefix	Suffix	add acid
simple anions HX (binary)	hydro-	-ic	acid
Polyatomic anions			
base	-ate	-ic	acid
(-1 oxygen)	-ite	-ous	acid

   Ex. HCl hydrochloric acid
   HI hydroiodic acid

SO_4^{2-}	sulfate	H_2SO_4	sulfuric acid
NO_3^-	nitrate	HNO_3	nitric acid
PO_4^{3-}	phosphate	H_3PO_4	phosphoric acid
SO_3^{2-}	sulfite	H_2SO_3	sulfurous acid
NO_2^-	nitrite	HNO_3	nitrous acid
PO_3^{3-}	phosphite	H_3PO_3	phosphorous acid

2. A) Bases- are ionic compounds that separate to give a metal ion and hydroxide ions. (-OH)

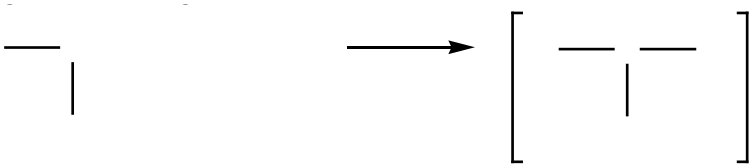
   $KOH \rightarrow K^+_{(aq)} + OH^-_{(aq)}$

   B) Nomenclature metal + hydroxide

   NaOH sodium hydroxide KOH potassium hydroxide

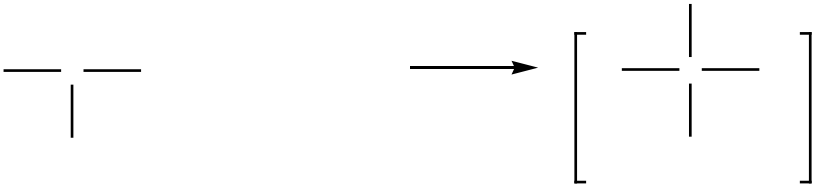
3. Further defining acids and bases. Bronsted-Lowry Acid and bases.

Acid- donates a proton (H^+)
base- daccepts a proton (H^+)



hydronium ion = H_3O^+

Base Acid Acid Base



$NaOH + H_2O \rightarrow H_2O + OH^-_{(aq)} + Na^+_{(aq)}$

4. **Strengths of Acids and Bases**

A) Strong Acids- ionize completely
Group 7A HCl, HBr, HI
Oxyacids HNO_3, H_2SO_4

$H_2O + HNO_3 \rightarrow H_3O^+ + NO_3^-_{(aq)}$
0% 100%

B) Weak Acids – Ionize very little

98% 2%

5. Bases

A) Strong bases- ionize completely to give metal ions and hydroxide (OH^-)

Group 1A and Ba are strong Bases.

$KOH \xrightarrow{H_2O} K^+_{(aq)} + OH^-$

B) Weak Bases – produce few hydroxide ions

Base Acid Acid Base

6. Acid- Base Neutralization

A) Net ionic

Mix: Acid + Base → water + salt neutralization

Strong acid, strong base always get $H^+ + OH^- \rightarrow H_2O$

Ex. 9-1: $HCl + NaOH \rightarrow NaCl_{(aq)} + H_2O$

$H^+ + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow Cl^-_{(aq)} + Na^+_{(aq)} + H_2O$

Net Ionic: $H^+ + OH^-_{(aq)} \rightarrow H_2O$ strong acid, strong base

Ex. 9-2: $2HBr + Ba(OH)_2 \rightarrow Ba(Br)_{2(aq)} + 2\,H_2O$

$2H^+ + 2Br^-_{(aq)} + Ba^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Br^-_{(aq)} + Ba^{2+}_{(aq)} + H_2O$

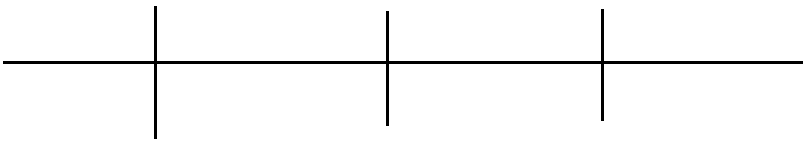
Net Ionic: $H^+ + OH^-_{(aq)} \rightarrow H_2O$ strong acid, strong base

B) acid- Base titration- neutralize an acid sample with a known base (use an indicator)

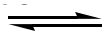
moles of acid (H^+) = moles of base (OH^-)
Ma x Va = Mb x Vb

Ex. 9-3: What is the molarity of a 5.00 mL sample of sulfuric acid if it is titrated with 40.0 mL of a 2.0 M sodium hydroxide solution.

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_{4(aq)} + 2H_2O$



6. Ion Product ion of water- water does ionize, but very little, 2 in 1 billion



base acid acid base

$[H_3O^+] = 1.0 \times 10^{-7}$ M $[OH^-] = 1.0 \times 10^{-7}$ M Brackets mean molarity

Ion product for water $K_w = [H_3O^+][OH^-] = [1.0 \times 10^{-7}][1.0 \times 10^{-7}] = 1.0 \times 10^{-14}$

neutral solution $[H_3O^+] = [OH^-]$
acidic solution $[H_3O^+] > [OH^-]$
basic solution $[H_3O^+] < [OH^-]$

Ex. 9-4. Are these solutions acidic or basic?

$[H_3O^+] = 1.0 \times 10^{-3}$ M $[H_3O^+] = 1.0 \times 10^{-10}$ M $[OH^-] = 1.0 \times 10^{-4}$ M
Acidic basic basic

K_w is constant regardless of acid or base added.

Ex. 9-5. What is the hydroxide ion concentration if the $[H_3O^+] = 1.0 \times 10^{-3}$ M?

$K_w = [H_3O^+][OH^-]$ $[OH^-] = K_w/[H_3O^+] = 1.0 \times 10^{-14}/ 1.0 \times 10^{-3} = 1.0 \times 10^{-11}$

7. pH scale, pOH

pH = 7 neutral	pOH = 7 neutral
pH > 7 basic	pOH > 7 acidic
pH < 7 acidic	pOH < 7 basic

show overhead of pH values

$pH = -\log[H_3O^+]$ $\quad$ $\text{invlog}(-pH) = [H_3O^+]$

$K_w = [H_3O^+][OH^-] = [1.0 \times 10^{-7}][1.0 \times 10^{-7}]$

$pH = -\log[1.0 \times 10^{-7}] = 7$

$pH + pOH = 14$

Ex. 9-6. A) If the $[H_3O^+] = 4.0\ \ 10^{-3}$ M what is the pH?

$pH = -\log[4.0\ \ 10^{-3}] = 2.4$

B) If the $[H_3O^+] = 8.2\ \ 10^{-8}$ M what is the pH?

$pH = -\log[8.2\ \ 10^{-9}] = 8.1$

C) What is the pOH for B?

$pH + pOH = 14$ $\quad$ $14 - 8.1 = 5.9$

Ex. 9-7. A) If the pH = 3.4 what is the H_3O^+ M?

$[H_3O^+] = \text{invlog}(-pH) = \text{invlog}(-3.4) = 4.0 \times 10^{-4}$

B) What is the $[OH^-]$ for A?

$pOH = 14 - 3.4 = 10.6$ $\quad [OH^-] = \text{invlog}(-pOH) = \text{invlog}(-10.6) = 2.5 \times 10^{-11}$

or $[OH^-] = K_w / [H_3O^+] = 1.0 \times 10^{-14} / 4.0\ \ 10^{-4} = 2.5 \times 10^{-11}$

8. Acid Rain-

Normal rain = pH = 6.2

SO_2 is produced when coal is burned with sulfur.

$S + O_2 \rightarrow SO_2$

$2SO_2 + O_2 \rightarrow 2SO_3$

$SO_3 + H_2O \rightarrow H_2SO_4$

marble $2H^+ + CaCO_3 \rightarrow CO_2 + H_2O + Ca^{2+}$

9. **Buffers**- resists changes in pH when small amounts of an acid or base is added.
   Blood is buffered.

Buffer solutions contain:

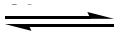
a) Weak acid, salt of the weak acid
   Weak base, salt of the weak base

Add salt of weak acid, replace H^+ with a cation such as group 1A

$NaHCO_3$ salt of weak acid

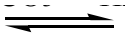
b) Add either base or acid to solution.

   1) Add base



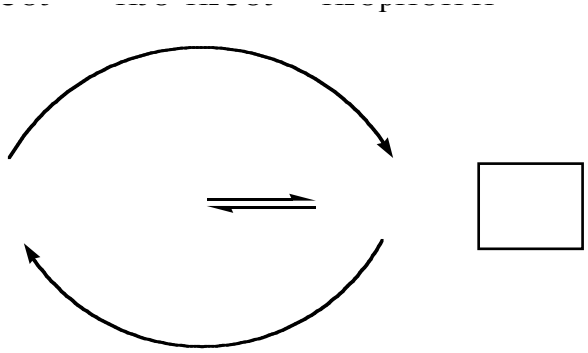
   no change in pH

   2) Add acid



   no change in Ph

Summary:



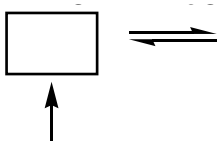
Ex. 9-8.

a) Write the ionization of HF in water.

b) What compound could be used to buffer this solution?

use the salt of the weak acid <u>NaF</u>

c) Write the net ionic equation for the addition of acid to the to the buffered solution.



d) Write the net ionic equation for the addition of base to the to the buffered solution.