Chapter 19 – Acids, Bases, and Salts

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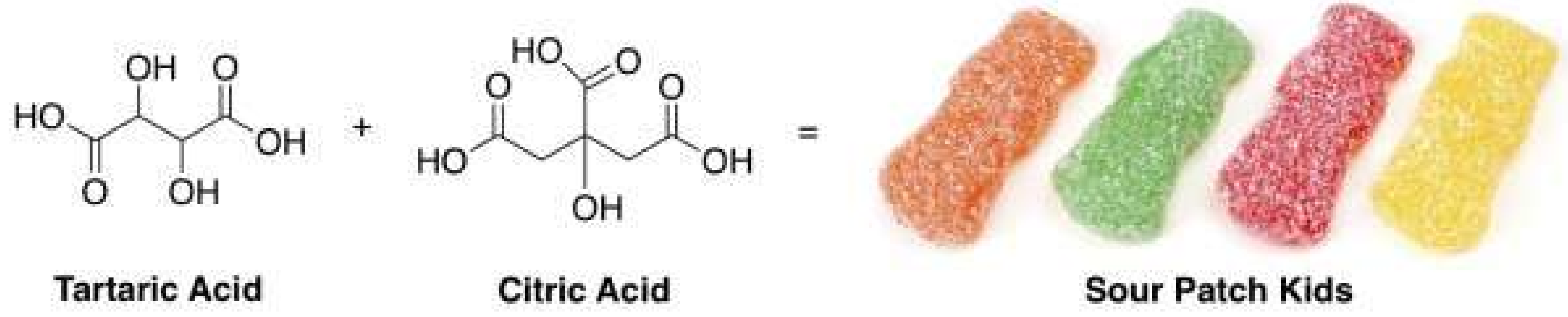
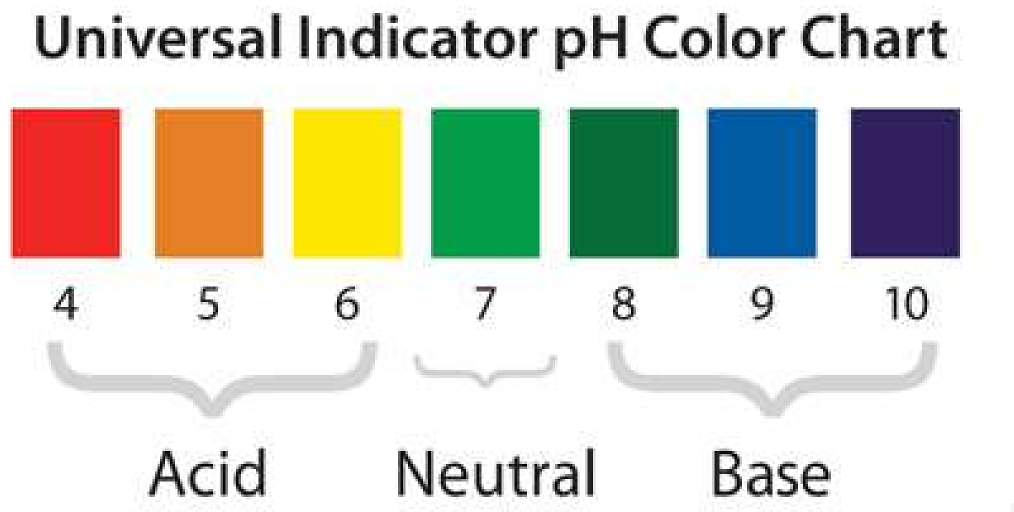
Section 19.1 – Acid-Base Theories

- Acids have a sour taste, change the color of an indicator, can be strong or weak electrolytes in aqueous solution, and react with metals.



Bases

- Bases taste bitter, change the color of an acid-base indicator, and can be strong or weak electrolytes in aqueous solution.



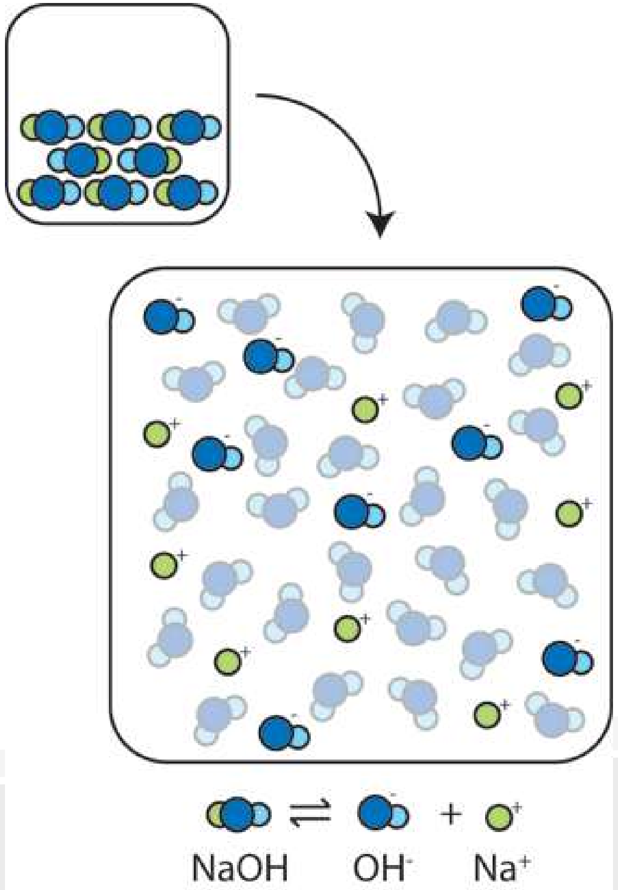
Arrhenius Acids

- Arrhenius acids are compounds that produce H^+ ions (H_3O^+) in a solution.
- A monoprotic acid produces 1 H^+ ion. Ex: HCl
- A diprotic acid produces 2 H^+ ions. Ex. H_2SO_4
- A triprotic acid produces 3 H^+ ions. Ex: H_3PO_4

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Arrhenius Bases

- Arrhenius bases are compounds that produce OH^- ions in solution. Ex: NaOH



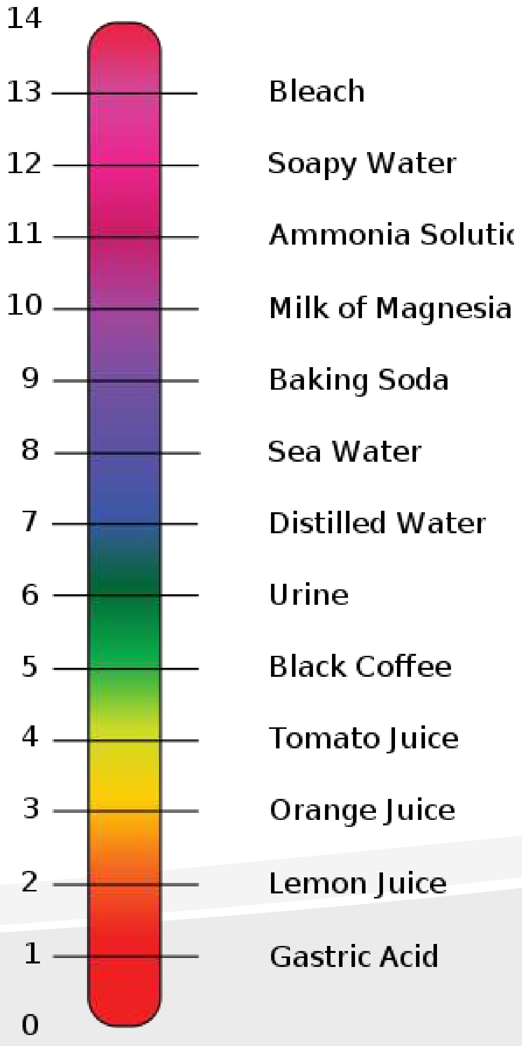
Bronsted-Lowry Acids and Bases

- H^+ ions are a proton.
- Bronsted-Lowry acids are proton (H^+) donors.

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

- Bronsted-Lowry bases are proton (H^+) acceptors.

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

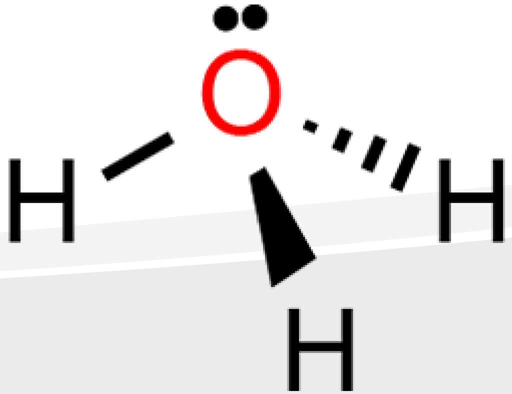


Conjugate Acid-Base Pair

- A conjugate acid is the particle formed when a base gains a hydrogen. (An acid created from a base)
- A conjugate base is the particle formed when a acid loses a hydrogen. (A base created from an acid)

$$HCl + H_2O \leftrightarrow H_3O^+ + Cl^-$$

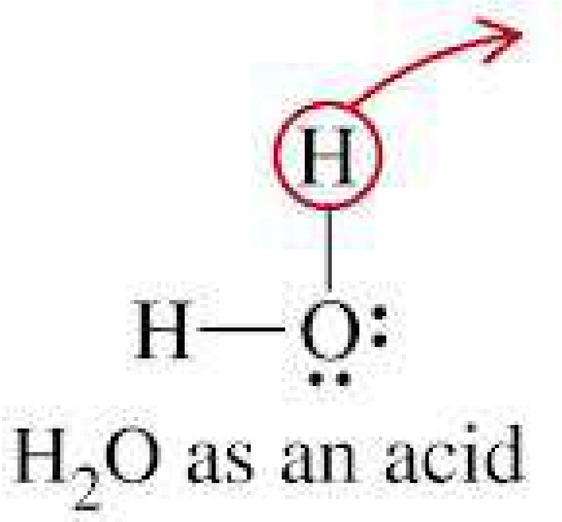
acid — base — conj. acid — conj. base



- H_3O^+ = hydronium ion

Amphoteric

- A substance that is <u>amphoteric</u> can act as an acid or a base.
- Ex: H_2O

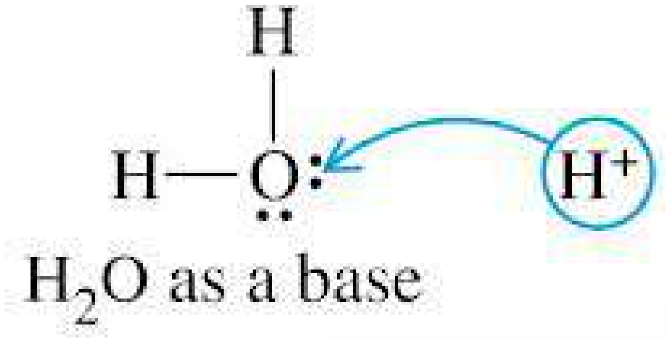


$HCl + H_2O \leftrightarrow H_3O^+ + Cl^-$

acid	base	conj. acid	conj. base

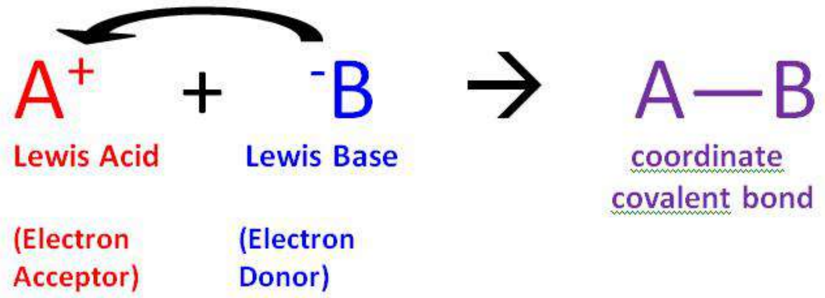
$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$

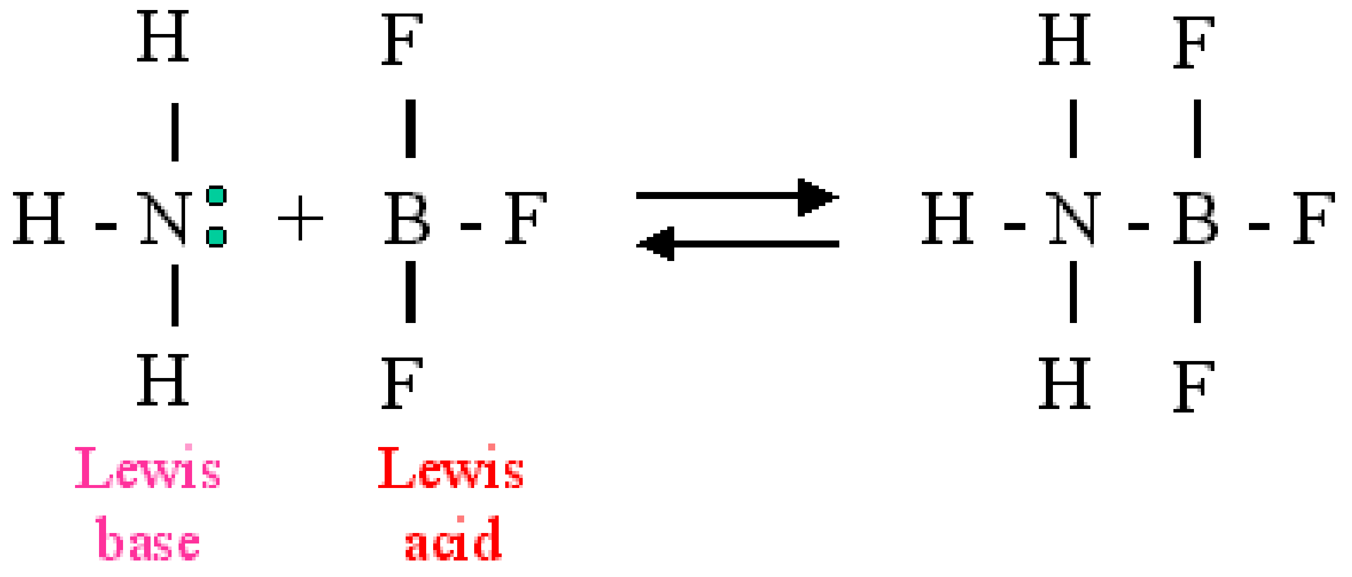
base	acid	conj. acid	conj. base



Lewis Acids and Bases

- <u>Lewis</u> acids are electron pair <u>acceptors</u>.
- <u>Lewis</u> bases are electron pair <u>donors</u>.





Section 19.1 Assessment

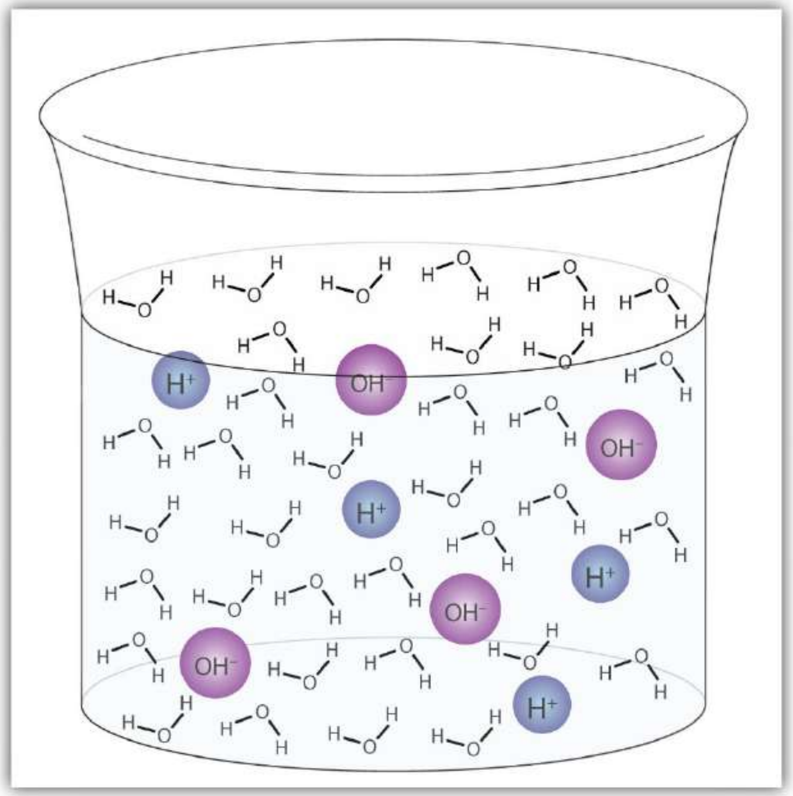
1. What are the properties of acids and bases?
2. How did Arrhenius define an acid and a base?
3. How are acids and bases defined by the Bronsted-Lowry theory?
4. What is the Lewis theory of acids and bases?
5. Identify the following acids as monoprotic, diprotic, or triprotic.

   a. H_2CO_3 b. H_3PO_4 c. HCl d. H_2SO_4

Section 19.2 – Hydrogen Ions and Acidity

- The reaction in which water molecules produce ions is called the self-ionization of water.

$$H_2O \leftrightarrow H^+ + OH^-$$

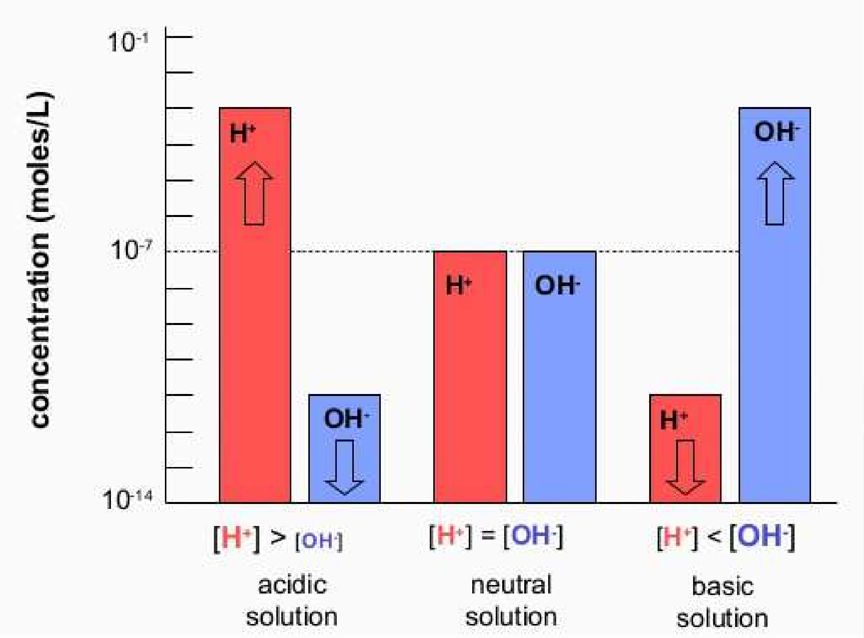


Ion Product Constant for Water

- In an aqueous solution, when $[H^+]$ increases, the $[OH^-]$ decreases and vice versa.
- However, the total product of the two concentrations is always 1×10^{-14}. This value is referred to a K_w (ion-product constant for water).

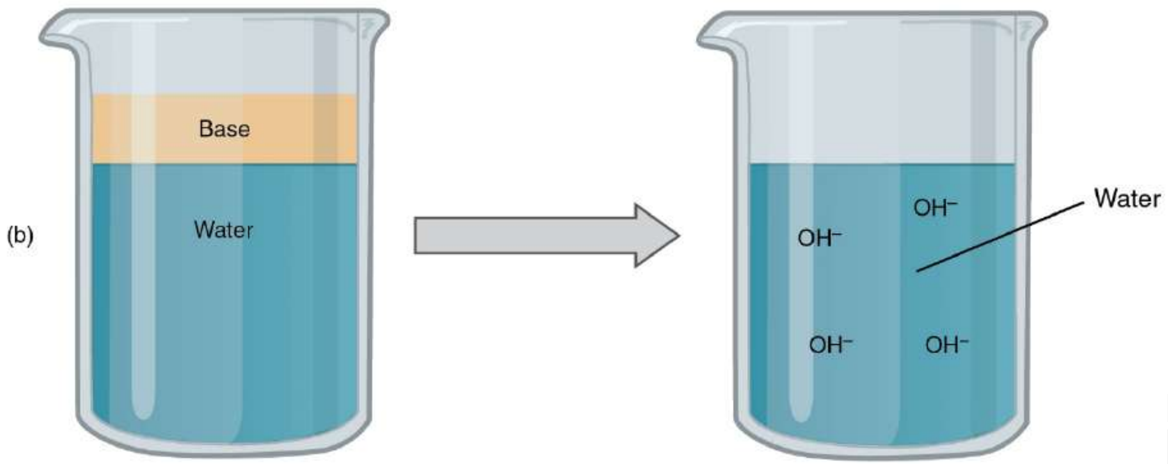
$$[H^+][OH^-] = 1 \times 10^{-14}$$

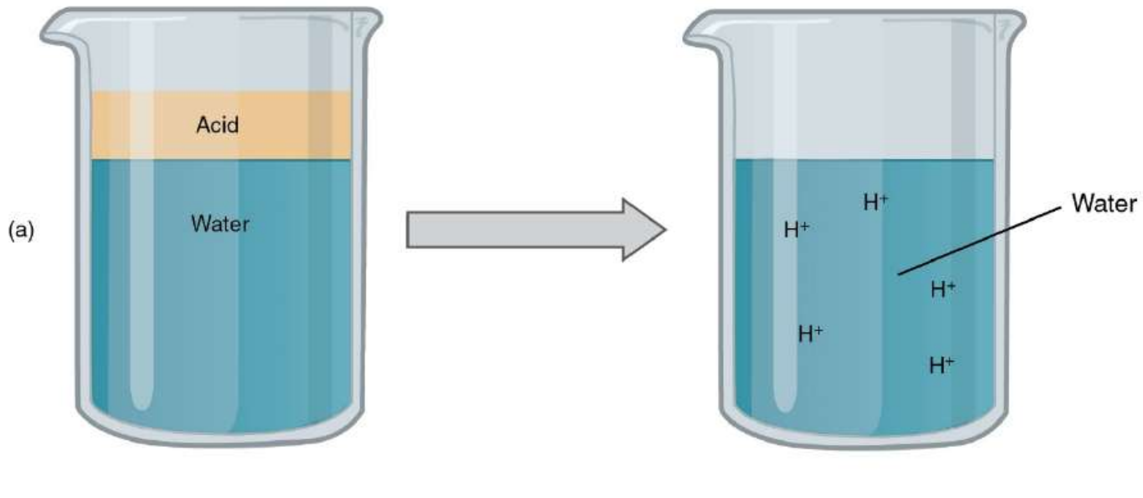
***K values have no units!!



Acidic, Basic, or Neutral

- In a neutral solution, $[H^+] = [OH^-] = 1 \times 10^{-7}M$
- In an acidic solution, the $[H^+]$ is larger than $[OH^-]$.
- In a basic solution, the $[OH^-]$ is larger than $[H^+]$.





Sample Problem

- If the $[H^+]$ in a coke is 1.0×10^{-5}M, what is the $[OH^-]$ and is the solution acidic, basic, or neutral?

$[OH^-] = 1.0 \times 10^{-9}$M

acidic

Practice Problems

1. Calculate the $[OH^-]$ of a solution that has an $[H^+] = 6.0 \times 10^{-10}M$. Is the solution acidic, basic, or neutral?

   $[OH^-] = 1.67 \times 10^{-5}M$
   basic

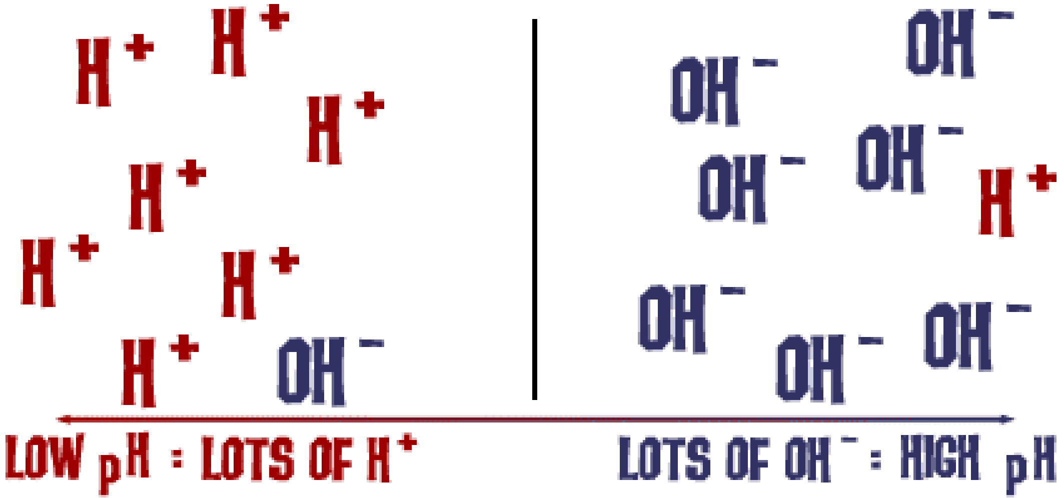
2. Calculate the $[H^+]$ of a solution that has a $[OH^-] = 3.0 \times 10^{-2}M$. Is the solution acidic, basic, or neutral?

   $[H^+] = 3.33 \times 10^{-13}M$
   basic

pH

- The <u>pH</u> of a solution is the negative log of the <u>hydrogen-ion</u> concentration.

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



- <u>Acidic</u> has a pH < 7
- <u>Neutral</u> has a pH = 7
- <u>Basic</u> has a pH > 7

***<u>pH</u> has <u>no units</u>!!

Sample Problem

- What is the pH of a solution with a hydrogen-ion concentration of 4.2×10^{-10}M and is the solution acidic, basic, or neutral?

9.38
Basic

Practice Problems

1. What is the pH of a solution that has an $[H^+] = 0.0015M$ and is the solution acidic, basic, or neutral?

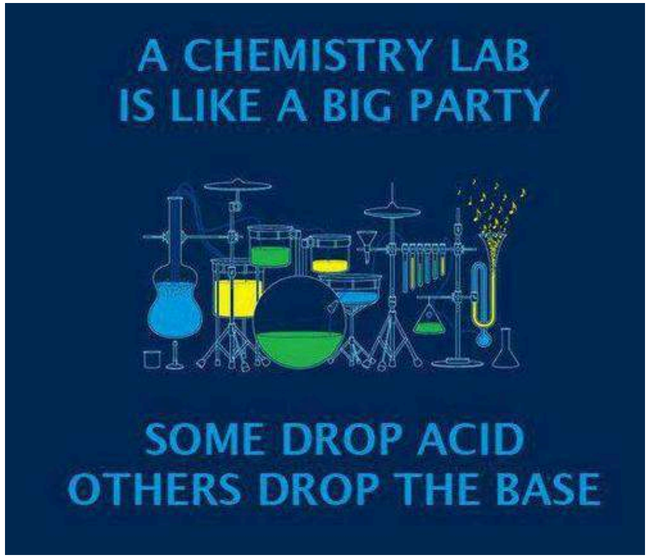
   2.82
   acidic

2. What is the pH value of a solution in which $[H^+] = 1.0 \times 10^{-12}M$ and is the solution acidic, basic, or neutral?

   12
   basic

pOH

- The pOH scale measures the OH^- concentration, so it is the opposite of the pH scale.



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

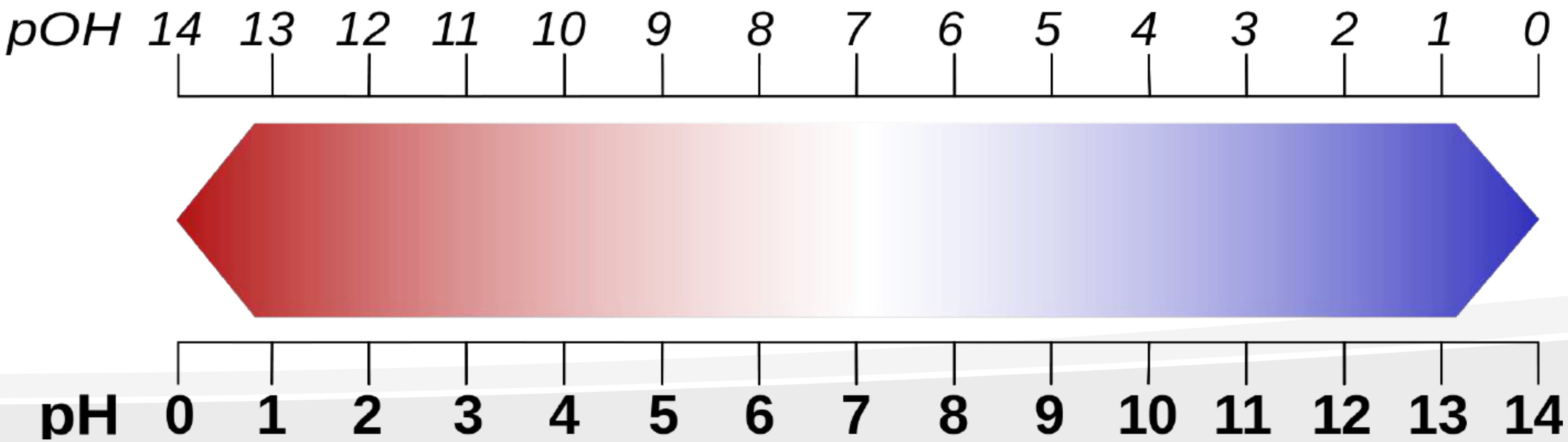
- Acidic has a pOH > 7
- Neutral has a pOH = 7
- Basic has a pOH < 7

***pOH has no units!!

pH vs. pOH

- The pOH scale is the reverse of the pH scale.

pH + pOH = 14



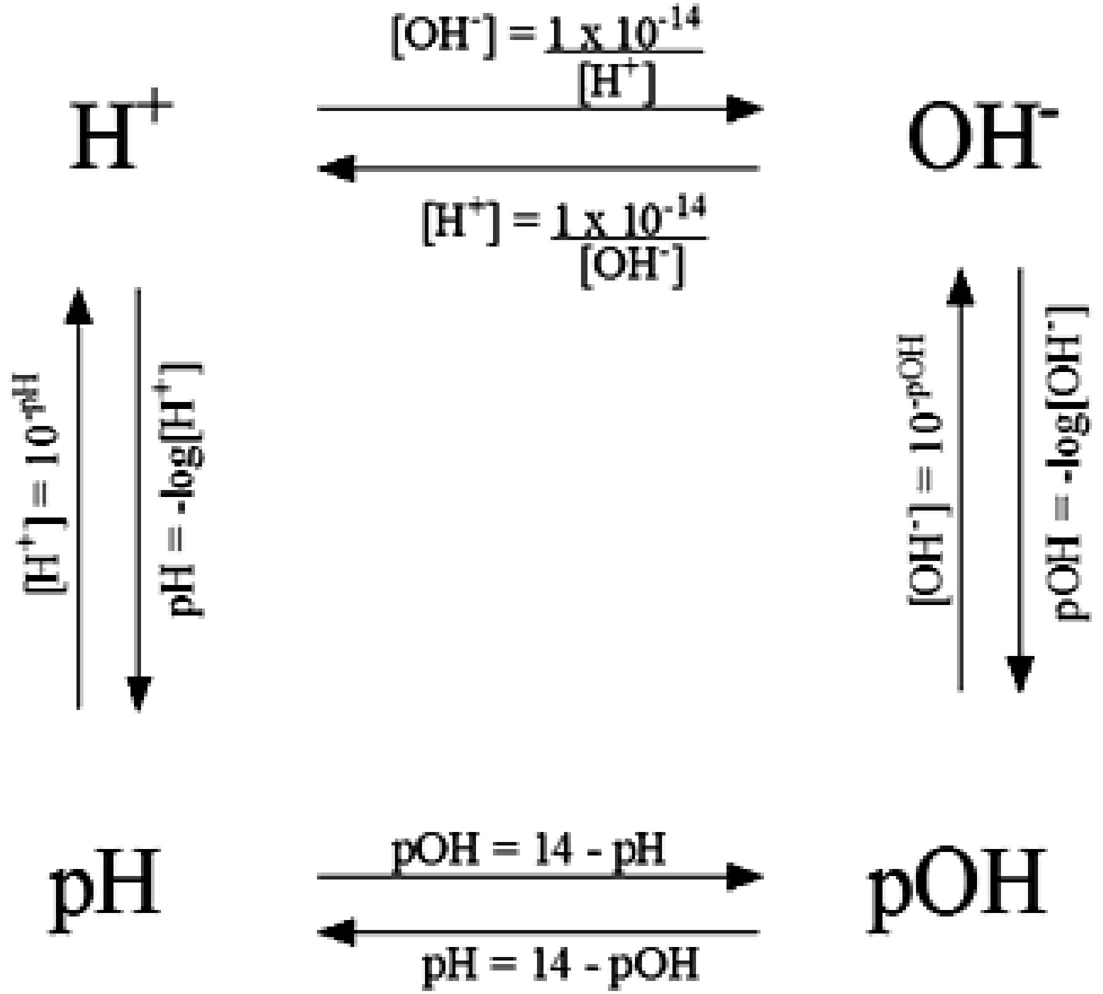
Calculating Concentration

- When going from the pH or pOH to concentration, you must <u>rearrange</u> the log formulas.

$$[H^+] = 10^{-pH}$$

$$[OH^-] = 10^{-pOH}$$

pH Square



[OH-] = 1 x 10-14 / [H+]
H+
OH-
[H+] = 1 x 10-14 / [OH-]
[H+] = 10-pH
pH = -log[H+]
[OH-] = 10-pOH
pOH = -log[OH-]
pOH = 14 - pH
pH
pOH
pH = 14 - pOH

Sample Problem

- The pH of an unknown solution is 6.35. What is the hydrogen-ion concentration and is the solution acidic, basic, or neutral?

$[H^+] = 4.5 \times 10^{-7} M$

acidic

Practice Problems

1. Calculate the pH of a solution with a pOH = 12.17 and is the solution acidic, basic, or neutral?

   1.83
   acidic

2. What is the pH of a solution if $[OH^-] = 4.0 \times 10^{-11}$M and is the solution acidic, basic, or neutral?

   3.60
   acidic

Acid-Base Indicators

- An acid-base indicator is a special chemical that changes color as the pH of a solution changes.

Section 19.2 Assessment

1. What is the relationship between $[H^+]$ and $[OH^-]$ in an aqueous solution?
2. What is true about the relative concentrations of hydrogen ions and hydroxide ions in each kind of solution?
   a. basic
   b. acidic
   c. neutral
3. Determine the pH of each solution.
   a. $[H^+] = 1 \times 10^{-6}M$
   b. $[OH^-] = 1 \times 10^{-2}M$
   c. $[H^+] = 0.00010M$
   d. $[OH^-] = 1 \times 10^{-11}M$

Section 19.2 Assessment

4. What are the hydroxide-ion concentrations for solutions with the following pH values?

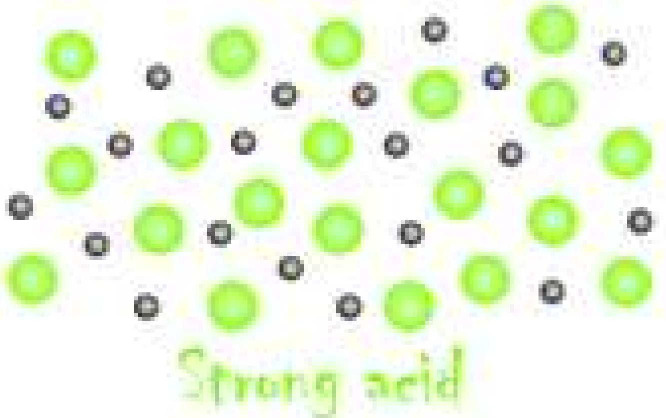
   a. 6.00

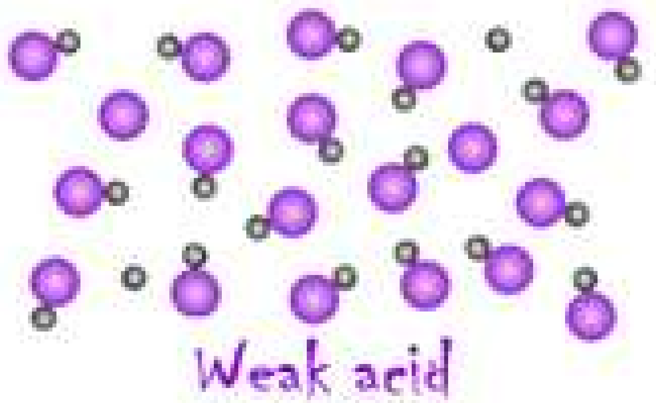
   b. 9.00

   c. 12.00

Section 19.3 – Strengths of Acids and Bases

- In general, strong acids completely dissociate in aqueous solution.
- Weak acids only slightly ionize in aqueous solution.
- Strong acids include HCl, HNO_3, and H_2SO_4.





Acid Dissociation Constant (K_a)

- The acid dissociation constant (K_a) is the ratio of the concentration of dissolved ions to the concentration of undissolved acid.

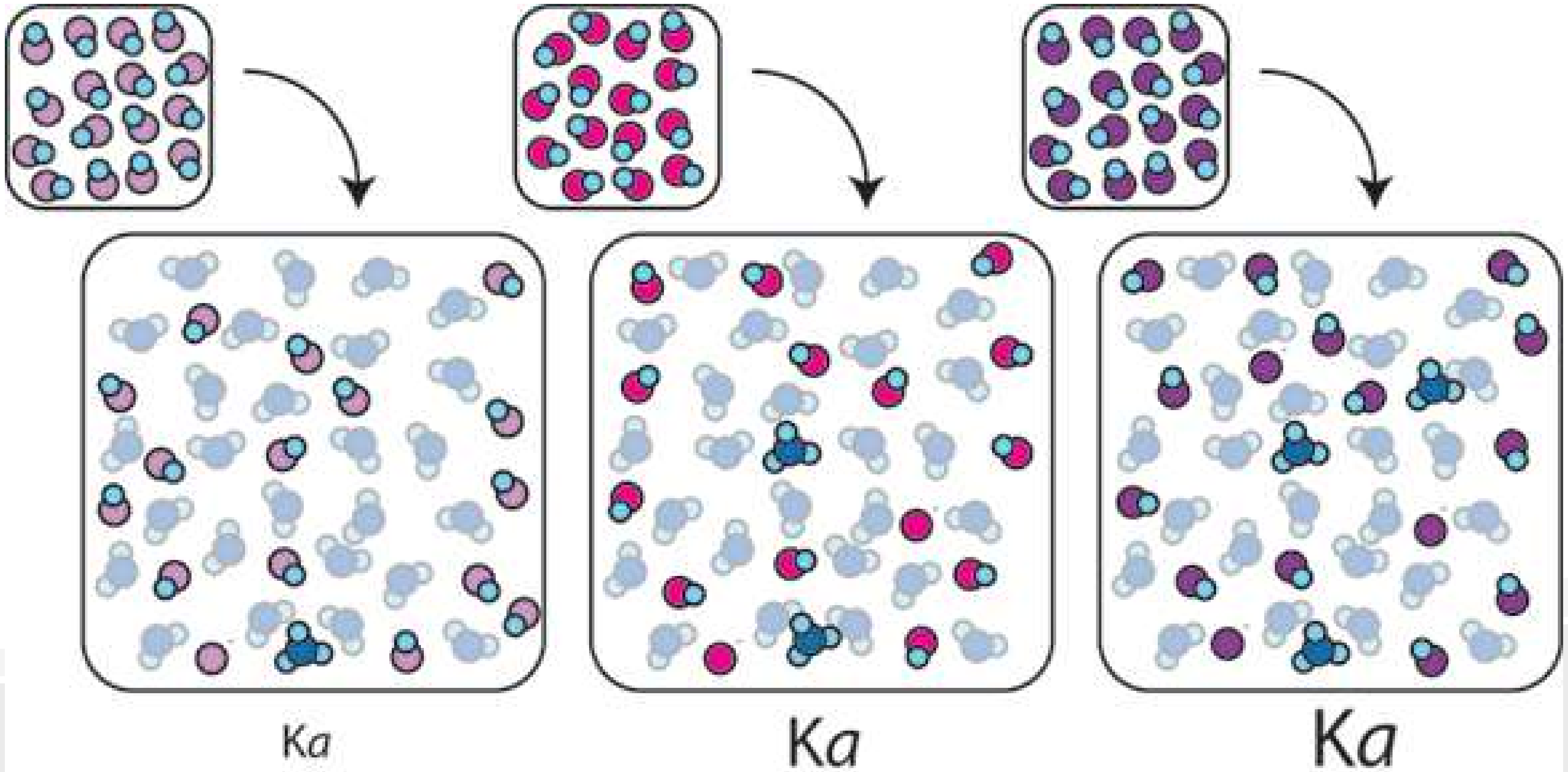
$$HNO_2 + H_2O \leftrightarrow H_3O^+ + NO_2^-$$

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

- A pure solid or liquid (H_2O) is not included in a K value.

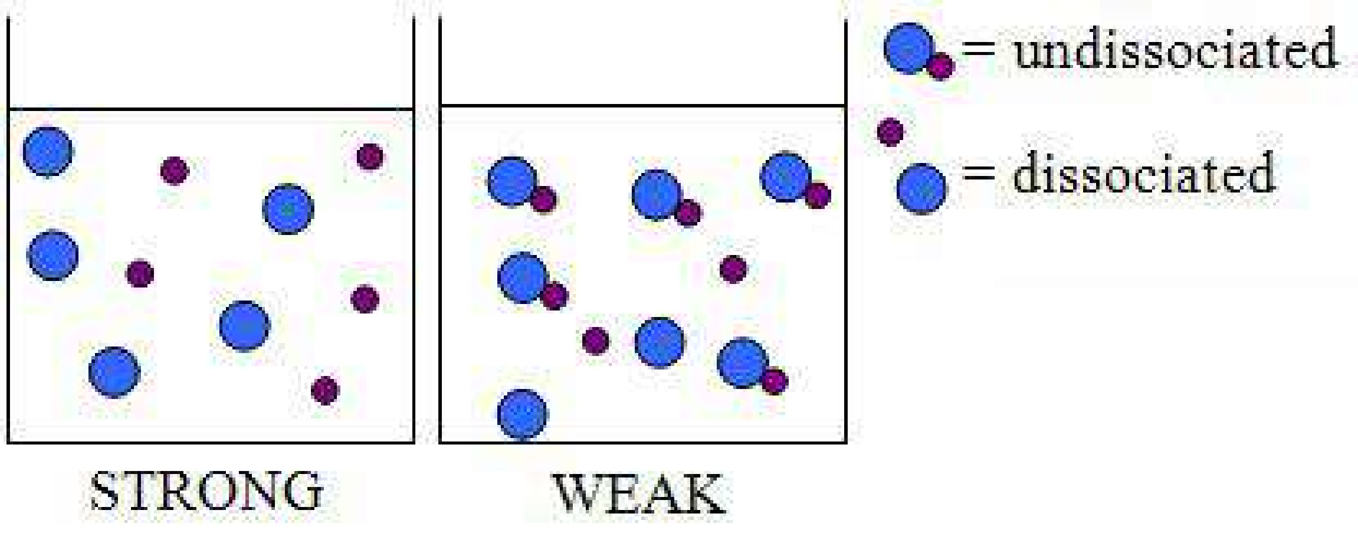
K_a

- The K_a value indicates the amount of ionized particles, so a weak acid has a small K_a and a strong acid has a large K_a.



Bases

- Strong bases fully ionize or dissociate in an aqueous solution.
- Weak bases partially ionize in an aqueous solution.
- Strong bases include NaOH, KOH, and LiOH.



Base Dissociation Constant (K_b)

- The base dissociation constant (K_b) is the ratio of the concentration of dissolved ions to the concentration of undissolved base.

$$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

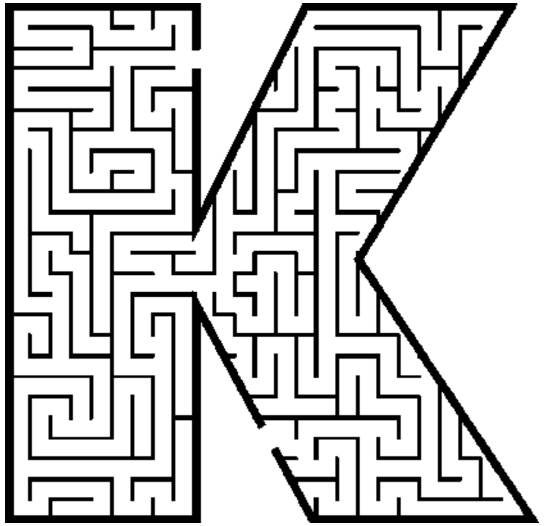
- The larger the K_b value, the stronger the base.

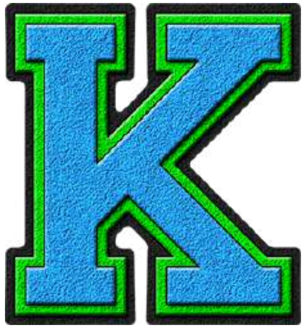
Generic K Equations

- The generic K_a formula:

$$HA \leftrightarrow H^+ + A^-$$

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





- The generic K_b formula:

$$B + H_2O \leftrightarrow BH^+ + OH^-$$

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

Sample Problem

- A 0.1000M solution of ethanoic acid is only partially ionized. From measurements of the pH of the solution, $[H^+]$ is determined to be 1.34×10^{-3}M. What is the acid dissociation constant (K_a) of ethanoic acid?

It doesn't matter whether or not you know what the formula for ethanoic acid is.

$$CH_3COOH + H_2O \leftrightarrow H_3O^+ + CH_3COO^-$$

Remember the generic formula

$$HA \leftrightarrow H^+ + A^-$$

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

Sample Problem Con't

Next you have to set up an ICE chart.

$$HA \leftrightarrow H^+ + A^-$$

	[HA]	$[H^+]$	$[A^-]$
Initial	0.1000M	0M	0M
Change	-1.34×10^{-3}M	$+1.34 \times 10^{-3}$M	$+1.34 \times 10^{-3}$M
Equilibrium	0.0987M	1.34×10^{-3}M	1.34×10^{-3}M

Sample Problem Con't

- You can only use EQUILBRIUM concentrations in a K equation.

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

$$K_a = \frac{[1.34 \times 10^{-3}M][1.34 \times 10^{-3}M]}{[0.0987M]} = \mathbf{1.82 \times 10^{-5}}$$

**K values have no units!!

Practice Problems

1. In a 0.1M solution of methanoic acid, $[H^+] = 4.2 \times 10^{-3}$M. Calculate the K_a of methanoic acid.

1.8×10^{-4}

2. In a 0.2M solution of a monoprotic weak acid, $[H^+] = 9.86 \times 10^{-4}$M. What is the K_a for this acid?

4.89×10^{-6}

Section 19.3 Assessment

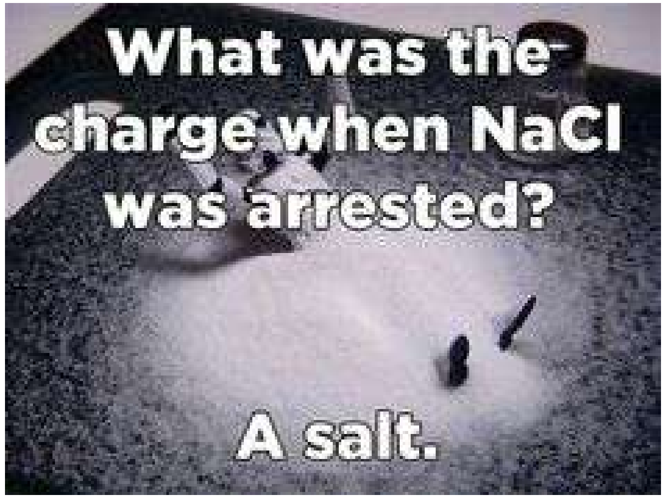
3. Compare a strong acid and a weak acid in terms of the acid dissociation constant.

Section 19.4 – Neutralization Reactions

- A neutralization reaction is a reaction between an acid and a base that forms water and a salt.
- Ex: $HCl + NaOH \rightarrow H_2O + NaCl$

  $H_2SO_4 + 2KOH \rightarrow 2H_2O + K_2SO_4$

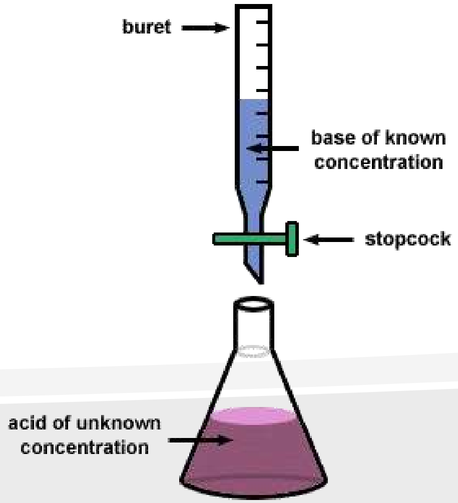
  acid base water salt

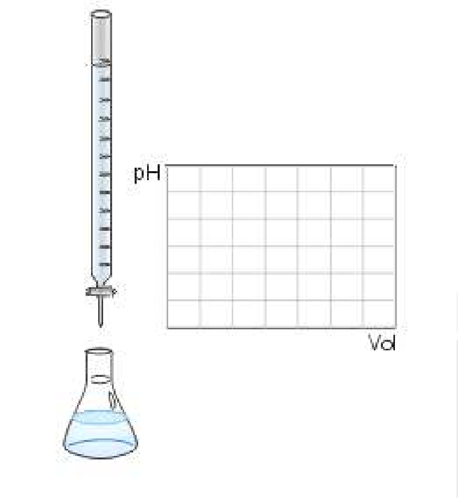


- A salt is a compound formed from the cation of a base and the anion of an acid.

Titration

- A titration is the use of a buret to add a measured amount of a known acid (or base) to a measured amount of an unknown base (or acid) until neutralization is achieved.
- The equivalence point of a titration is when the number of moles of hydrogen ions equals the number of moles of hydroxide ions.





Sample Problem

- How many moles of sulfuric acid are required to neutralize 0.50 mol of sodium hydroxide?

$$H_2SO_4 + 2NaOH \rightarrow 2H_2O + Na_2SO_4$$

$$0.50 \text{ mol NaOH} \times \frac{1 \text{ mol } H_2SO_4}{2 \text{ mol NaOH}} = 0.25 \text{ mol } H_2SO_4$$

Practice Problems

1. How many moles of potassium hydroxide are needed to completely neutralize 1.56 mol of phosphoric acid?

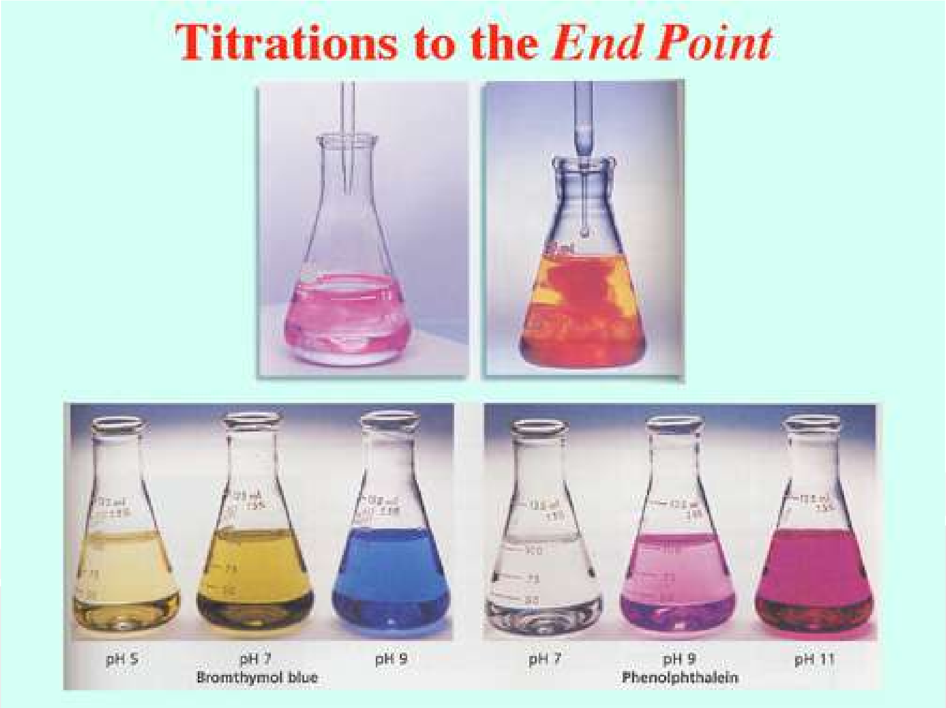
   4.68 mol KOH

2. How many moles of sodium hydroxide are required to neutralize 0.20 mole of nitric acid?

   0.20 mol NaOH

Titration

- The <u>end point</u> of a titration is the point at which the indicator changes <u>color</u>.
- In the best titrations, the <u>end point</u> corresponds to the <u>equivalence point</u>.



Sample Problem

- A 25mL solution of H_2SO_4 is completely neutralized by 18mL of 1.0M NaOH. What is the concentration of the H_2SO_4 solution?

$H_2SO_4 + 2NaOH \rightarrow 2H_2O + Na_2SO_4$

M = mol/L so mol = M x L

Mol NaOH = 1.0M x 0.018L = 0.018 mol NaOH

$$0.018 \text{ mol NaOH} \times \frac{1 \text{ mol } H_2SO_4}{2 \text{ mol NaOH}} = 9 \times 10^{-3} \text{ mol } H_2SO_4$$

M = mol/L so $M = 9 \times 10^{-3}\text{mol}/0.025\text{L} =$ **$0.36M\ H_2SO_4$**

Practice Problems

1. How many milliliters of 0.45M HCl will neutralize 25.0mL of 1.00M KOH?

56mL HCl

2. What is the molarity of H_3PO_4 if 15.0mL is completely neutralized by 38.5mL of 0.150M NaOH?

0.129M H_3PO_4

Section 19.4 Assessment

1. What are the products of a reaction between an acid and a base?
2. How many moles of HCl are required to neutralize aqueous solutions of these bases?

   a. 2 mol NH_3

   b. 0.1 mol $Ca(OH)_2$
3. Write complete balanced equations for the following acid-base reactions.

   a. $H_2SO_4 + KOH \rightarrow$

   b. $H_3PO_4 + Ca(OH)_2 \rightarrow$

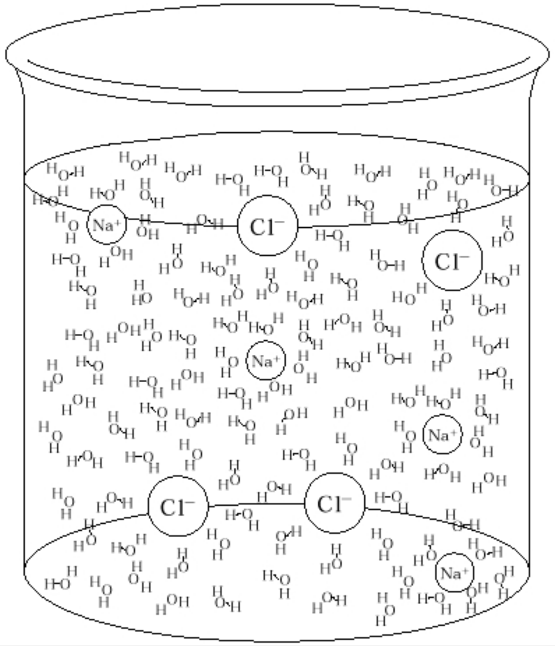
   c. $HNO_3 + Mg(OH)_2 \rightarrow$

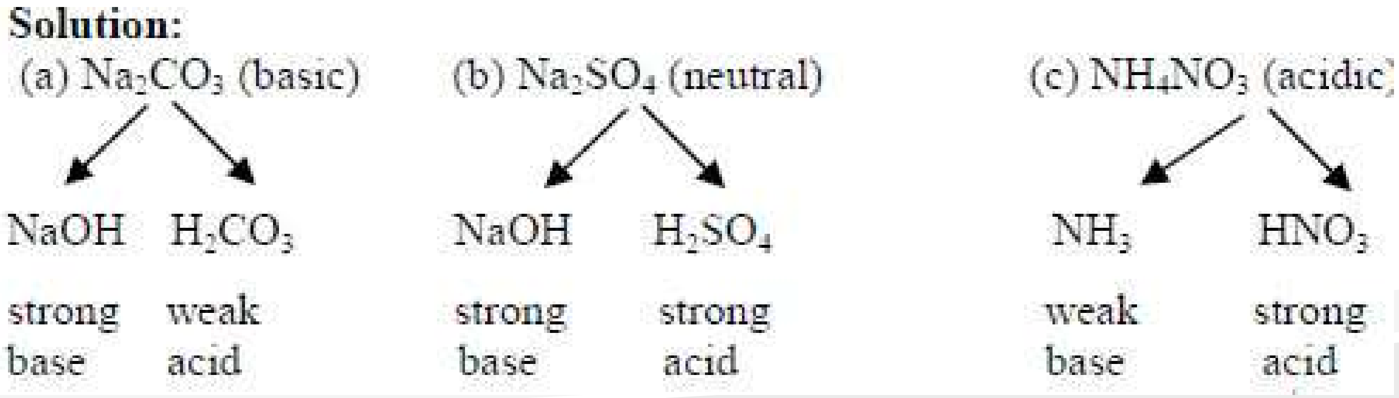
Section 19.5 – Salts in Solutions

- Remember: A salt is a compound formed from the cation of a base and the anion of an acid.
- A salt solution can be acidic, basic, or neutral.

Salt Solutions

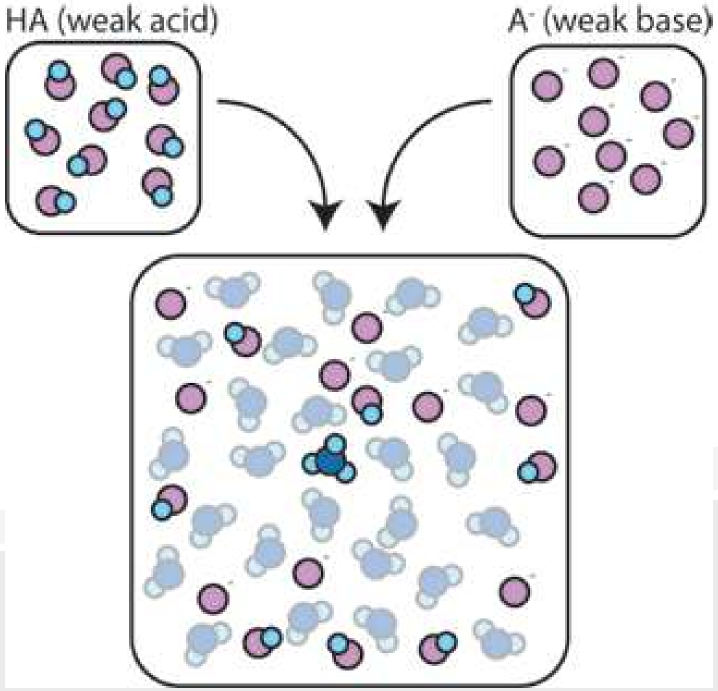
- Strong Acid + Strong Base = Neutral Solution
- Strong Acid + Weak Base = Acidic Solution
- Weak Acid + Strong Base = Basic Solution

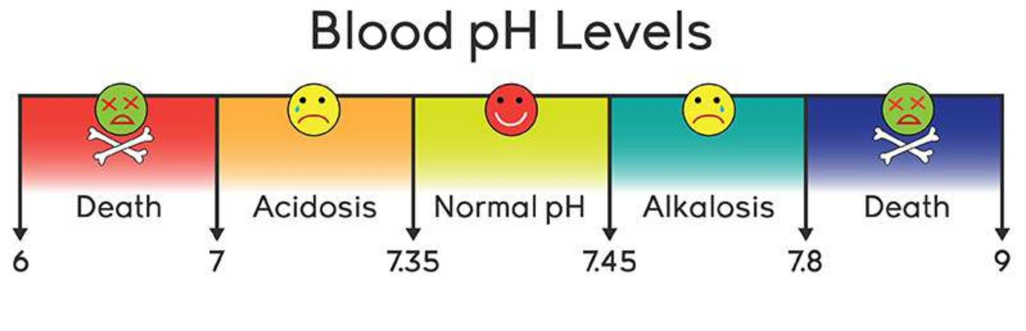




Buffers

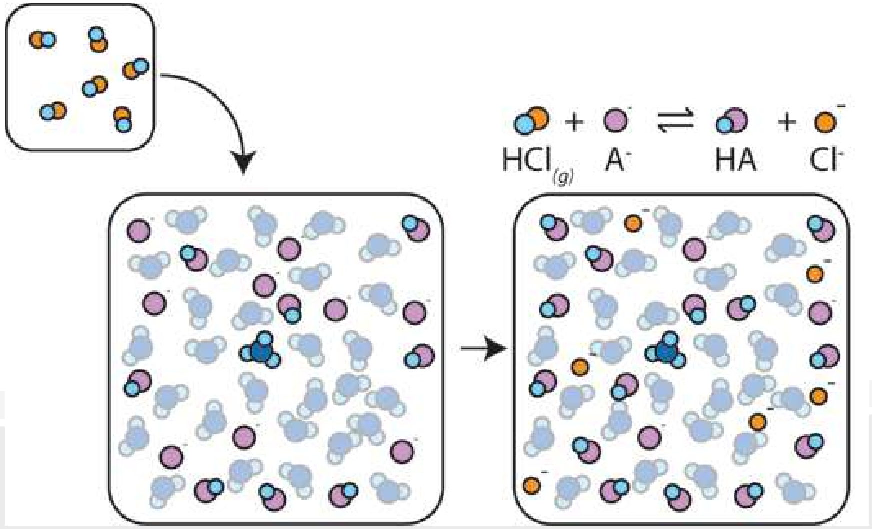
- A buffer is a solution in which the pH remains relatively constant when small amounts of acid or base are added.
- A buffer is made from a weak acid and its conjugate base or a weak base and its conjugate acid.
- A common buffer that you have is your blood.

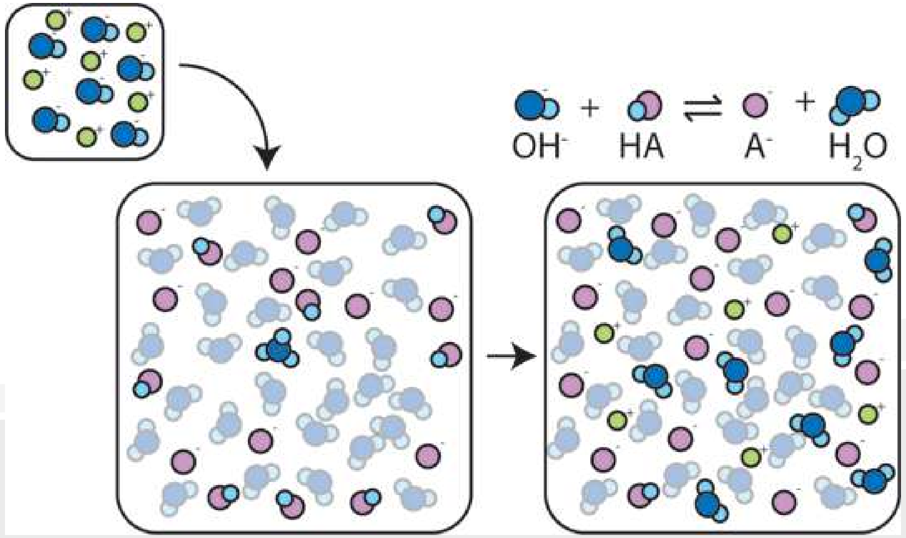




Buffer

- Since a buffer contains both an acidic and basic component, it can neutralize acid or base that is added.
- The buffer capacity is the amount of acid or base that can be added to a buffer solution before a significant change in pH occurs.





Section 19.5 Assessment

1. What substances are combined to make a buffer?
2. Which of these salts would form an acidic aqueous solution?

   a. $KC_2H_3O_2$

   b. LiCl

   c. $NaHCO_3$

   d. $(NH_4)_2SO_4$

THE END