Chapter 19 – Acids, Bases, and Salts

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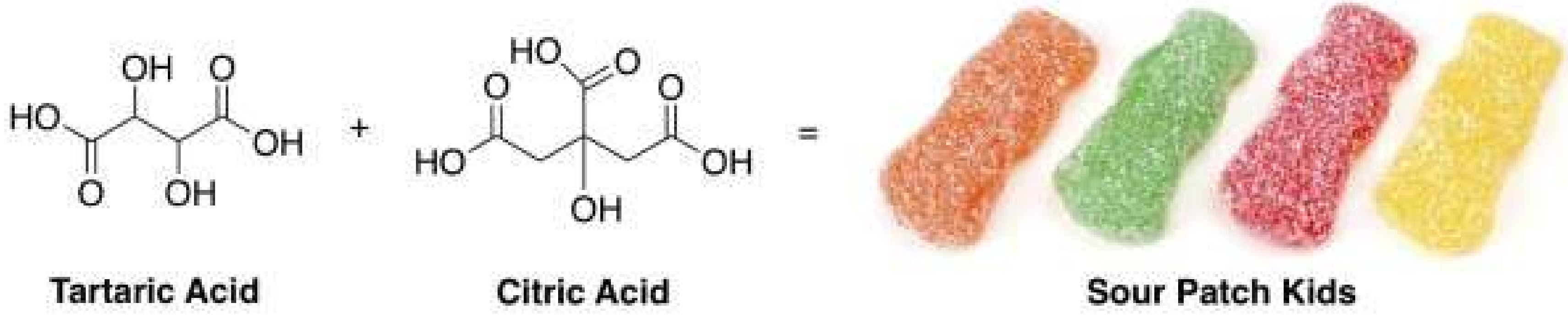
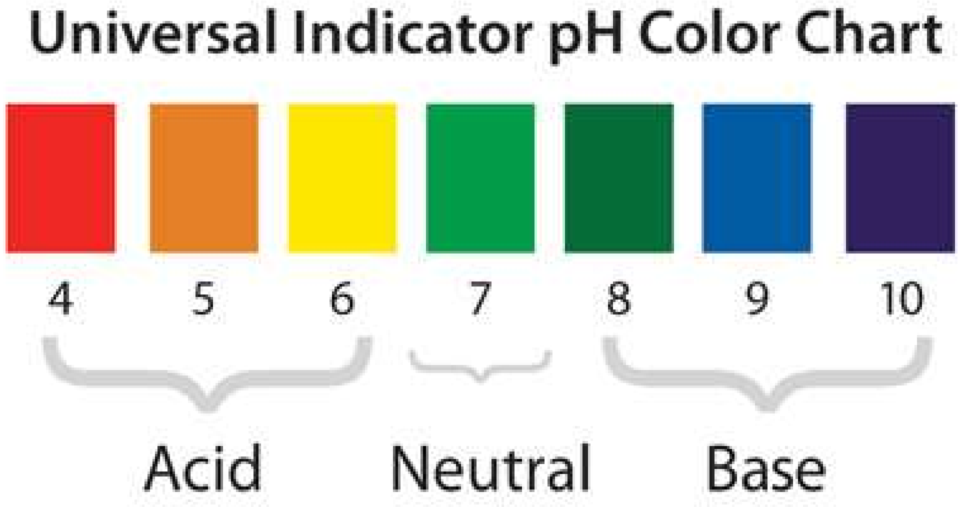
Section 19.1 – Acid-Base Theories

- 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.



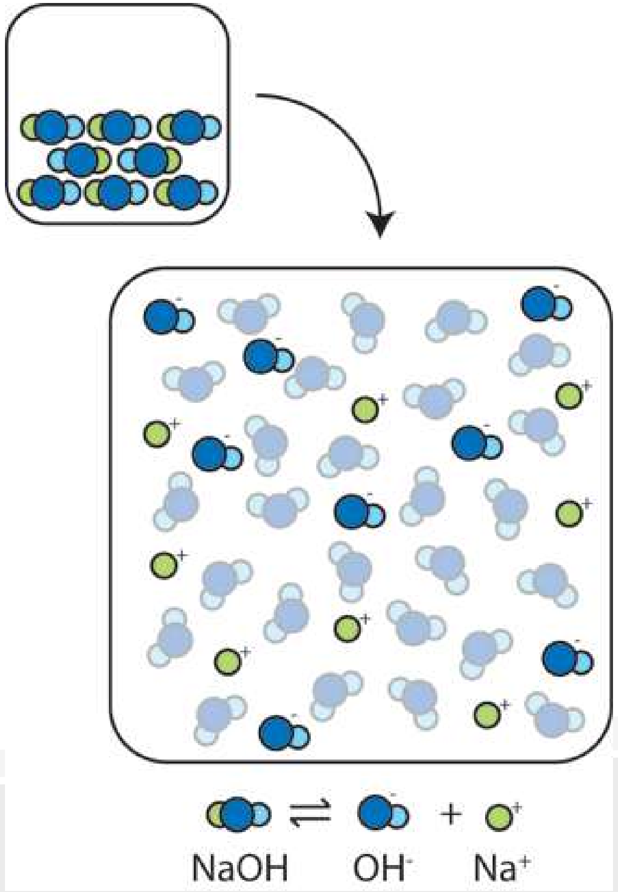
Arrhenius Acids

- 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



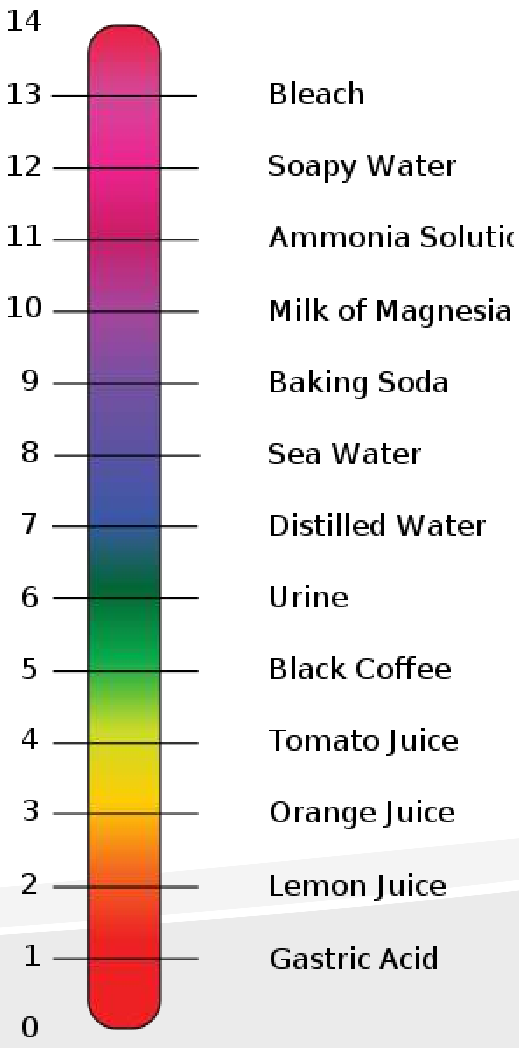
Bronsted-Lowry Acids and Bases

- 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.

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

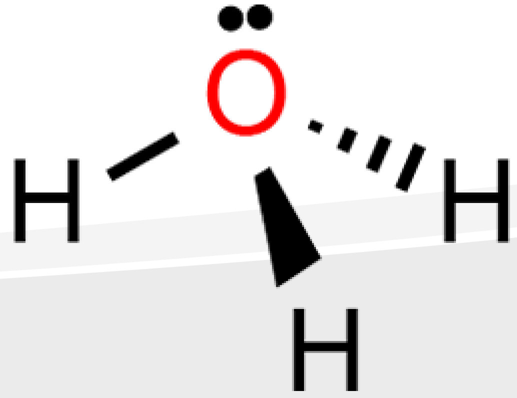


Conjugate Acid-Base Pair

- 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)

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

HCl	H_2O	H_3O^+	Cl^-
acid	base	conj. acid	conj. base



- H_3O^+ = hydronium ion

Sample Problem

- Write the conjugate base for the following acids:
- HCl
- HSO_4^-
- H_2O

Practice Problem

- Write the conjugate acid for the following bases:

- NH_3

- SO_4^{2-}

- HPO_4^{2-}

Practice Problem

- Which of the following are conjugate acid-base pairs?

a. HNO_3, NO_3^-

b. H_3PO_4, PO_4^{3-}

c. H_2SO_4, HSO_3^-

d. H_3PO_3, $H_2PO_3^{2-}$

e. HCN, H_2O

Sample Problem

- Label the acid, base, conjugate acid, and conjugate base in the following reaction:

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

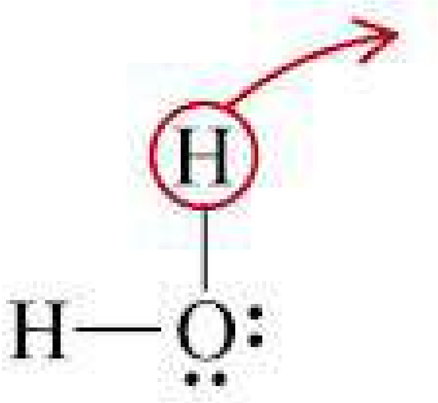
Practice Problem

- Label the acid, base, conjugate acid, and conjugate base in the following reaction:

$$OH^- + HOCl \leftrightarrow OCl^- + H_2O$$

Amphoteric

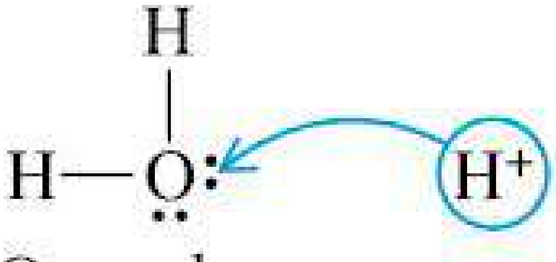
- A substance that is amphoteric can act as an acid or a base.
- Ex: H_2O



H_2O as an acid

HCl	+	H_2O	$\leftrightarrow$	H_3O^+	+	Cl^-
acid		base		conj. acid		conj. base

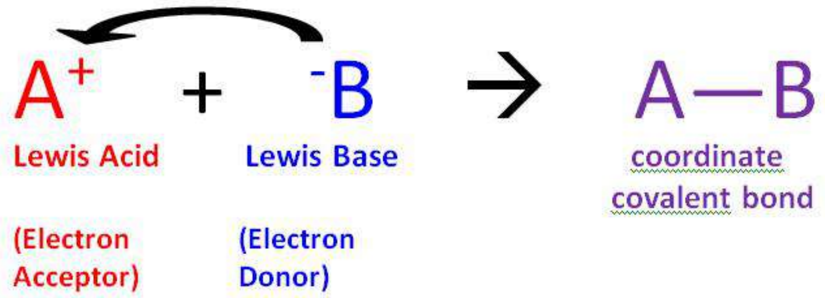
NH_3	+	H_2O	$\leftrightarrow$	NH_4^+	+	OH^-
base		acid		conj. acid		conj. base

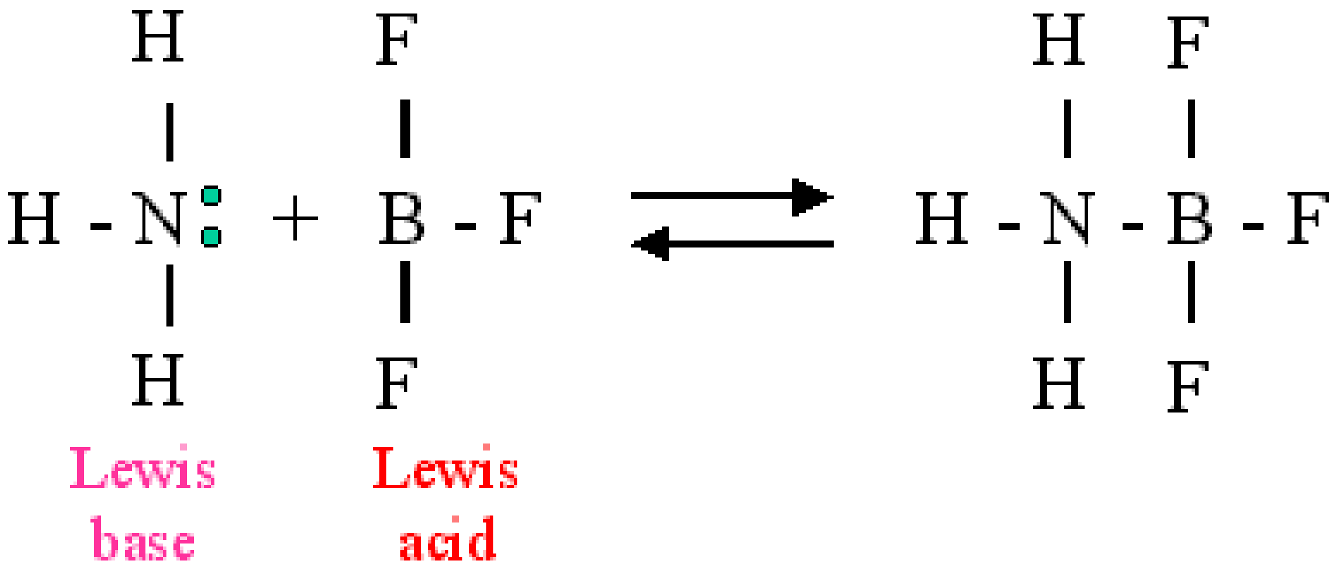


H_2O as a base

Lewis Acids and Bases

- Lewis acids are electron pair acceptors.
- Lewis bases are electron pair donors.





Section 19.1 Assessment

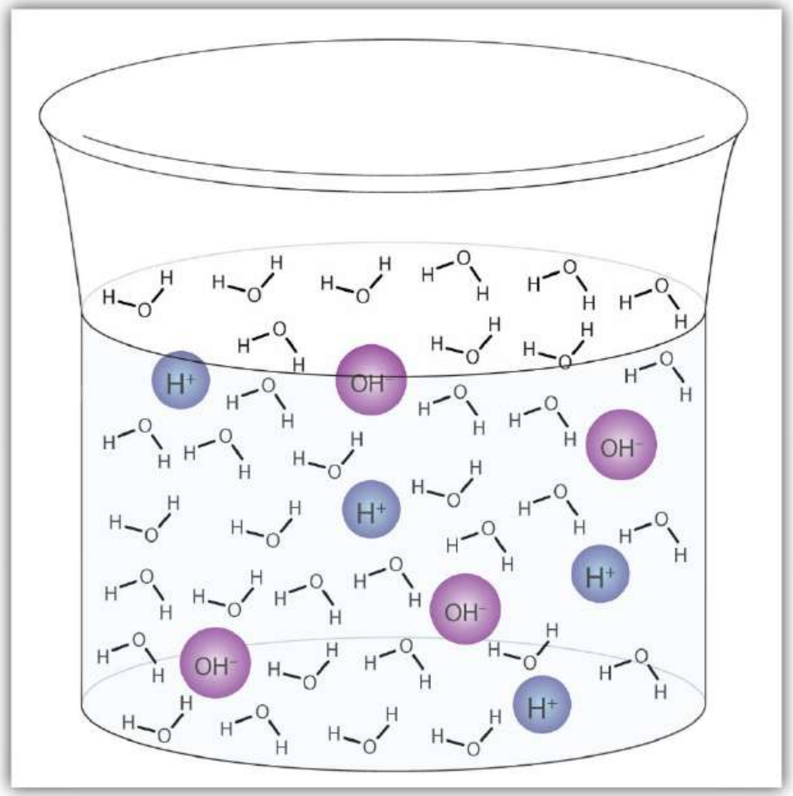
1. What are the properties of acids and bases?
2. How did Arrhenius define an acid and a base?
3. How are acids and bases defined by the Bronsted-Lowry theory?
4. What is the Lewis theory of acids and bases?
5. Identify the following acids as monoprotic, diprotic, or triprotic.

   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.

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

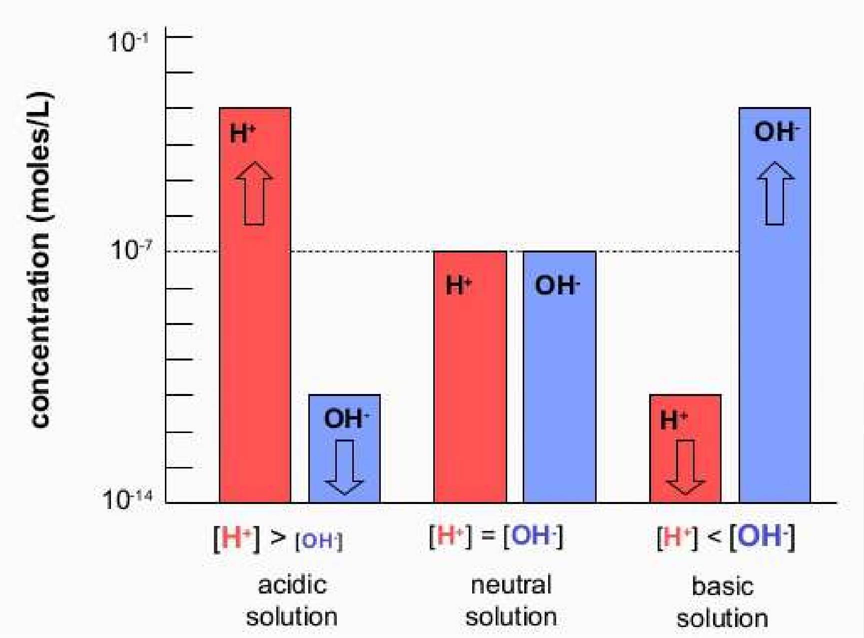


Ion Product Constant for Water

- 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).

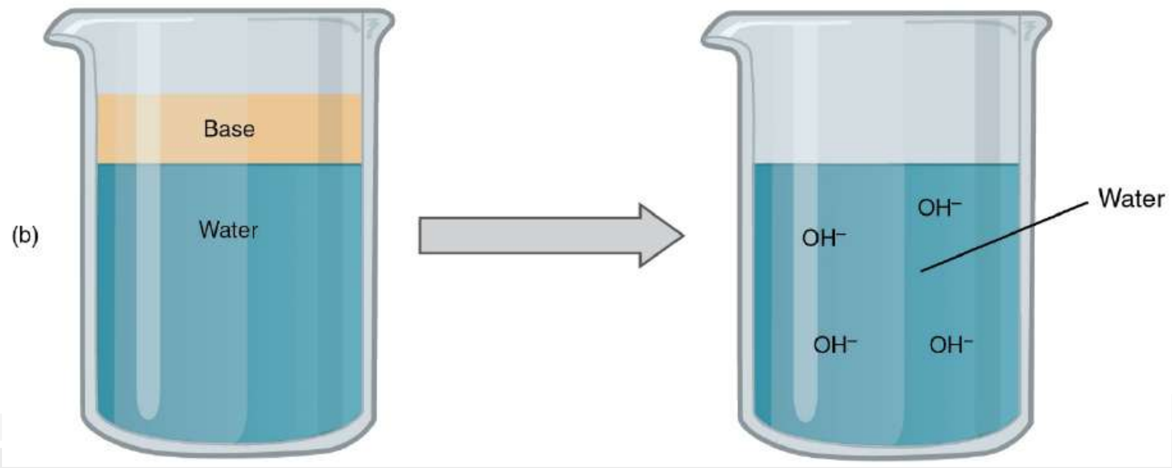
$$[H^+][OH^-] = 1 \times 10^{-14}$$

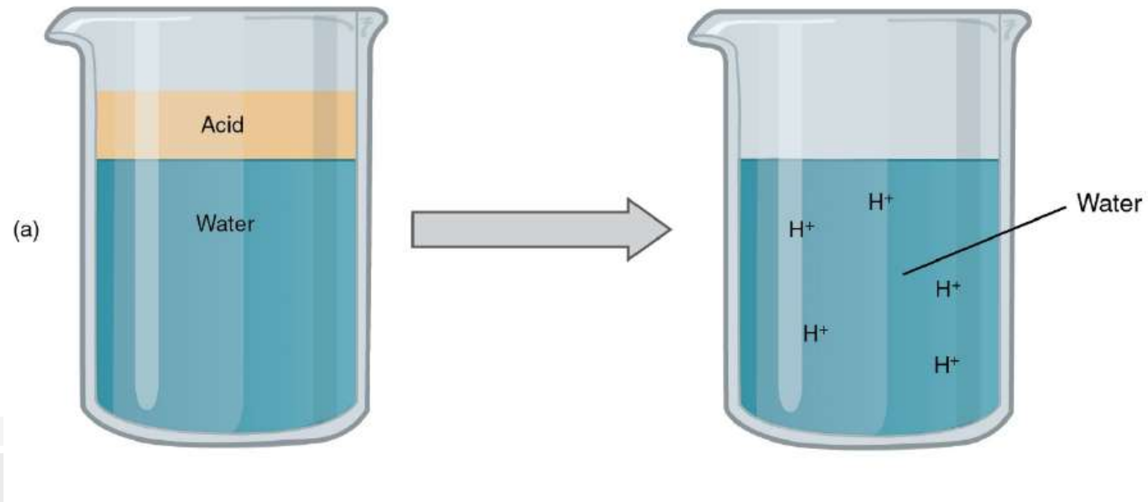
***K values have no units!!



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?

Practice Problems

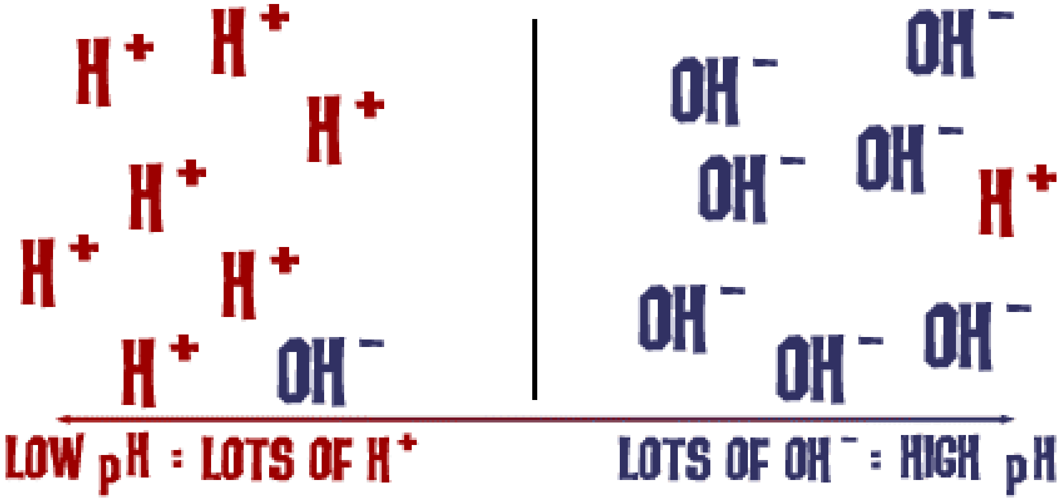
1. Calculate the $[OH^-]$ of a solution that has an $[H^+] = 6.0 \times 10^{-10}$M. Is the solution acidic, basic, or neutral?

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

pH

- The pH of a solution is the negative log of the hydrogen-ion concentration.

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



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

***pH has no units!!

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?

Practice Problems

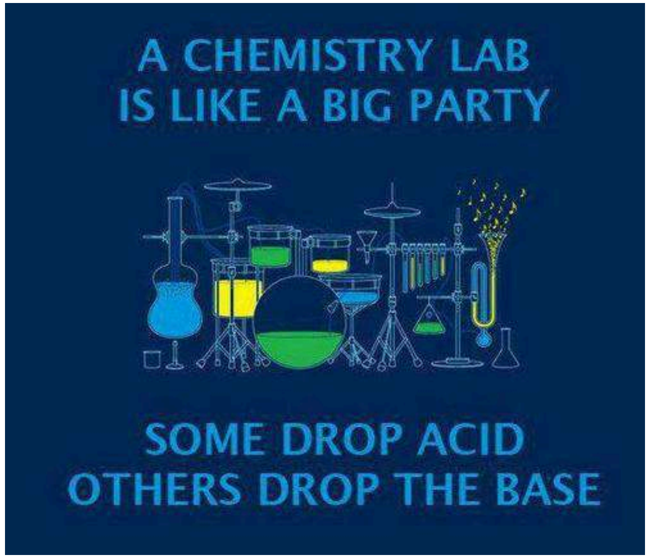
1. What is the pH of a solution that has an $[H^+] = 0.0015M$ and is the solution acidic, basic, or neutral?

2. 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?

pOH

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

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



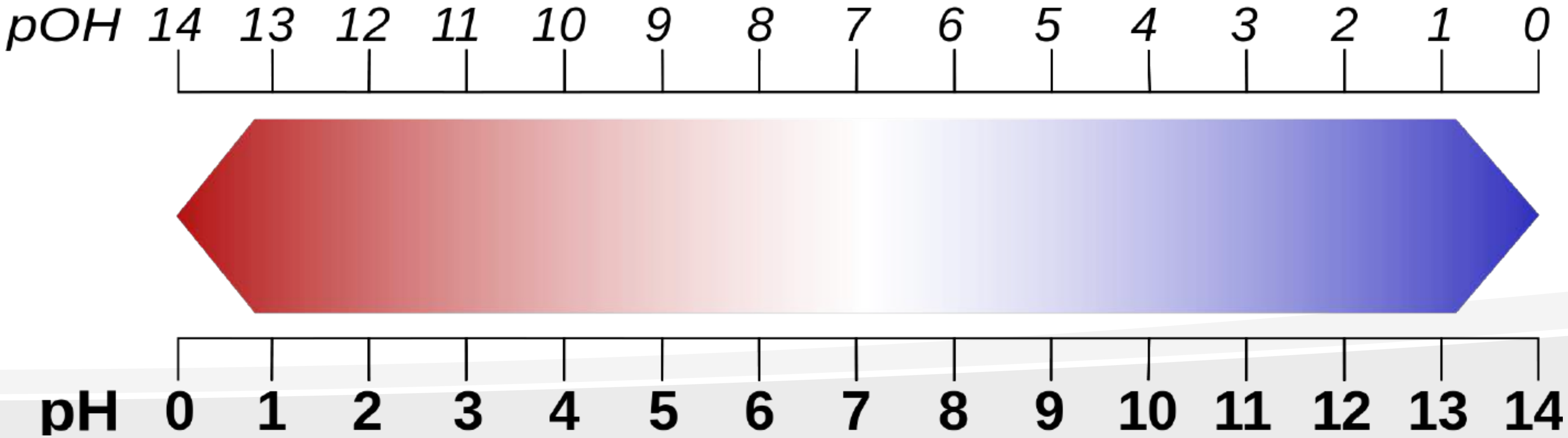
- Acidic has a pOH > 7
- Neutral has a pOH = 7
- Basic has a pOH < 7

***pOH has no units!!

pH vs. pOH

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

$$pH + pOH = 14$$



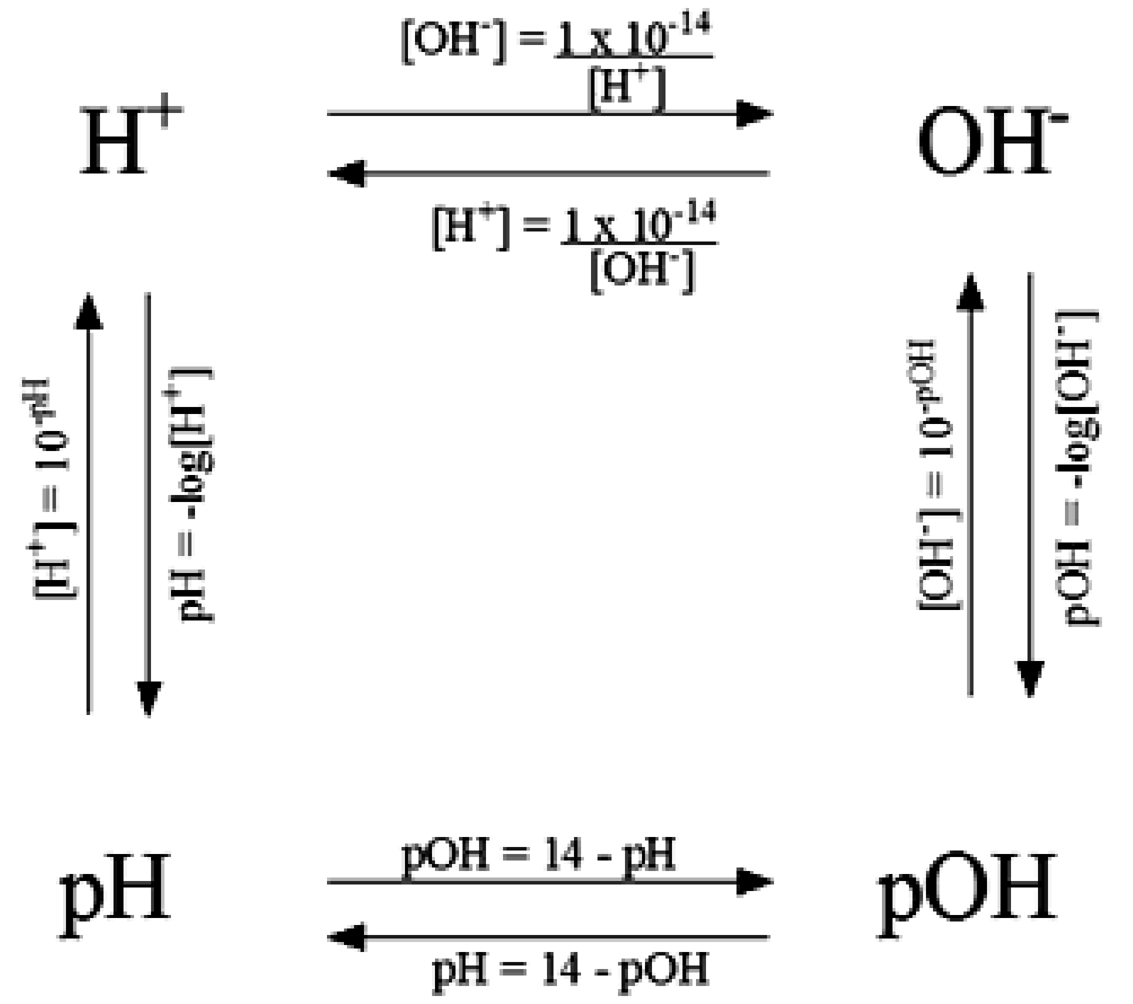
Calculating Concentration

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

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

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

pH Square



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

Practice Problems

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

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

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

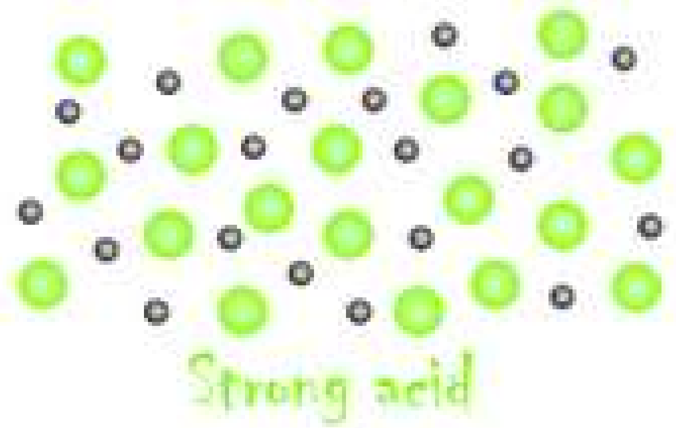
1. What is the relationship between $[H^+]$ and $[OH^-]$ in an aqueous solution?
2. What is true about the relative concentrations of hydrogen ions and hydroxide ions in each kind of solution?
   a. basic
   b. acidic
   c. neutral
3. Determine the pH of each solution.
   a. $[H^+] = 1 \times 10^{-6}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

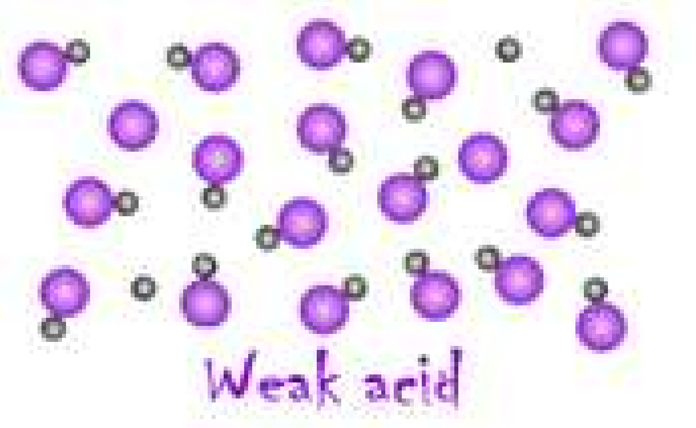
Section 19.2 Assessment

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, H_2SO_4, HBr, HI, $HClO_3$, and $HClO_4$.





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.

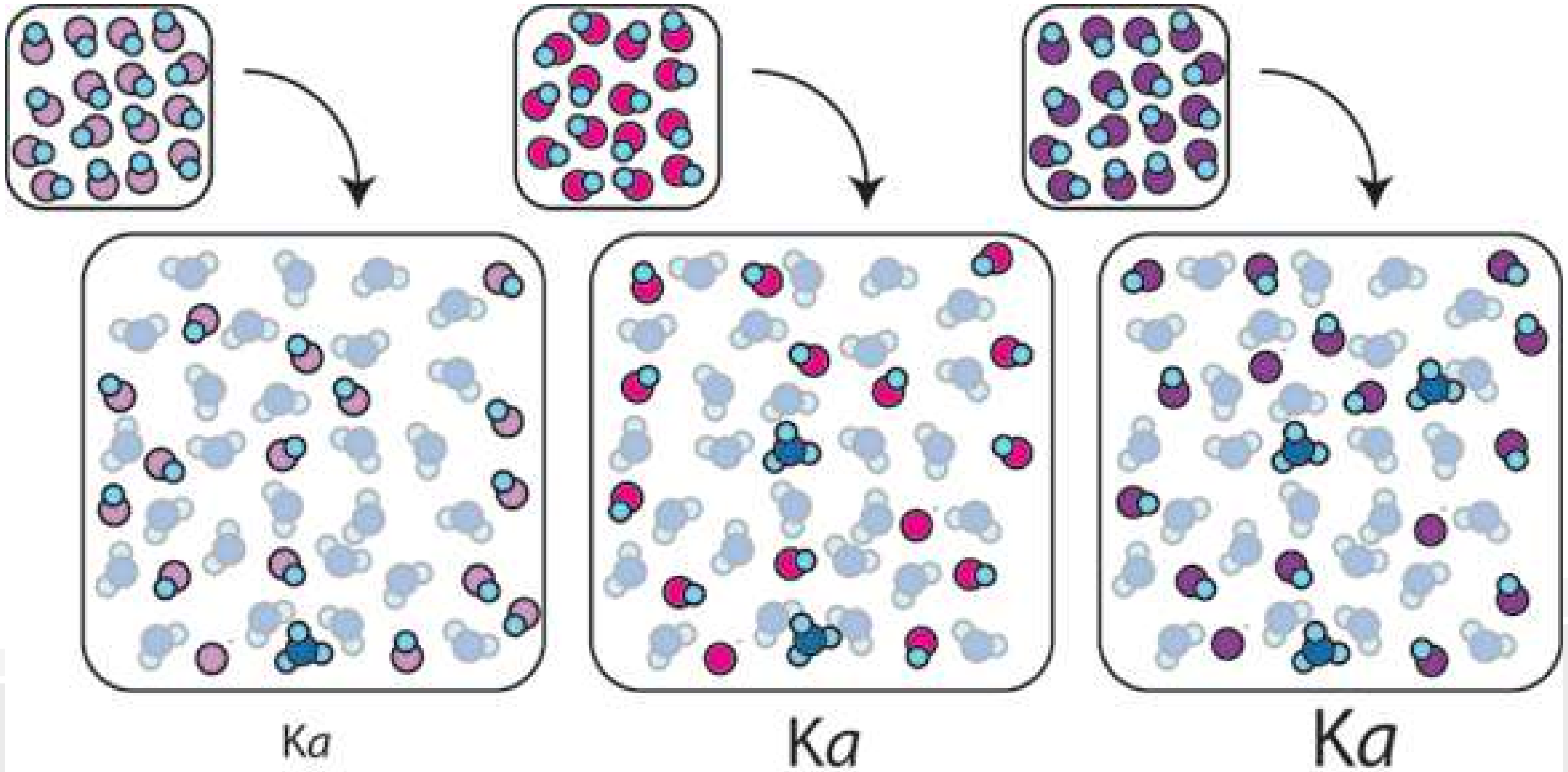
$$HNO_2 + H_2O \leftrightarrow H_3O^+ + NO_2^-$$

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

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

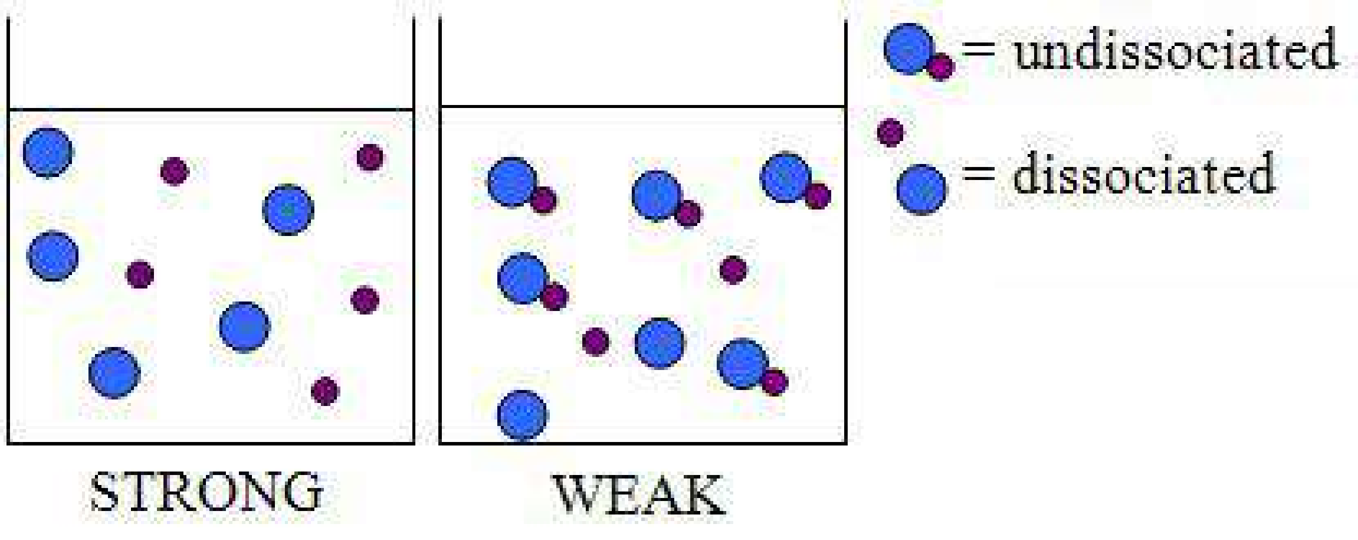
K_a

- 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, LiOH, RbOH, CsOH, $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$.
- The most common weak base is NH_3.



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.

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

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

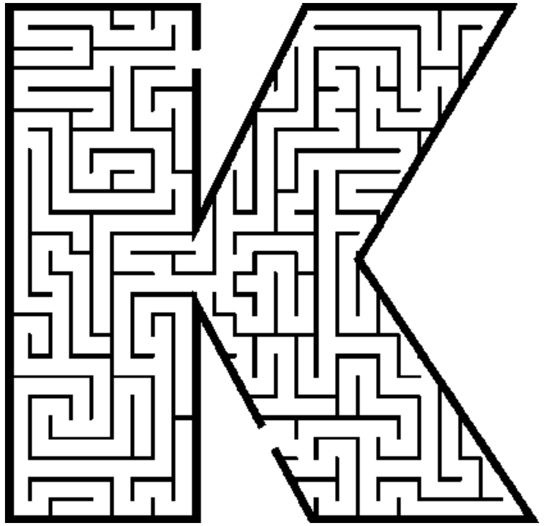
- The larger the K_b value, the stronger the base.

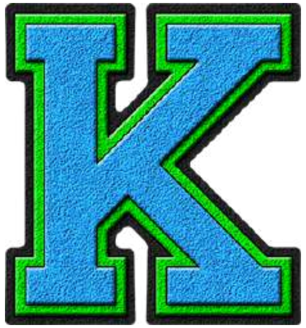
Generic K Equations

- 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?

Sample Problem Con’t

Sample Problem Con't

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.

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?

Section 19.3 Assessment

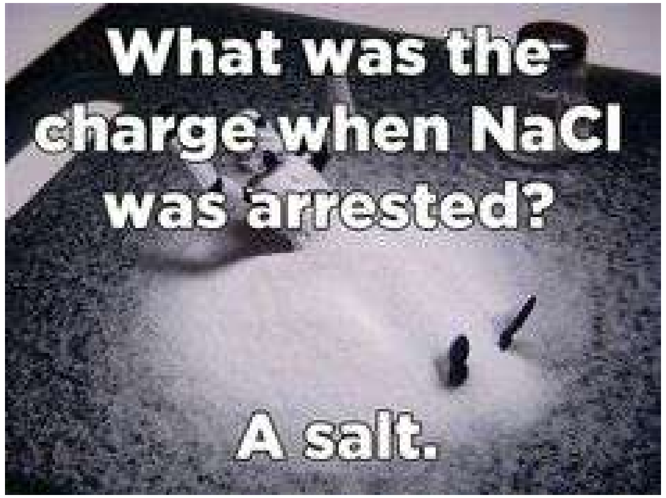
3. Compare a strong acid and a weak acid in terms of the acid dissociation constant.

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.

Sample Exercise

a. Write a balanced equation for the reaction between aqueous solutions of acetic acid (CH_3COOH) and barium hydroxide.

b. Write the net ionic equation for this reaction.

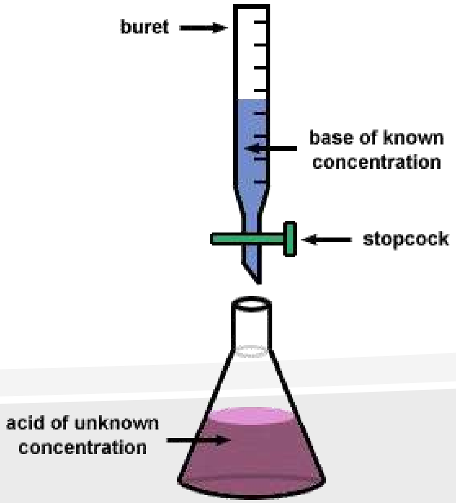
Practice Exercise

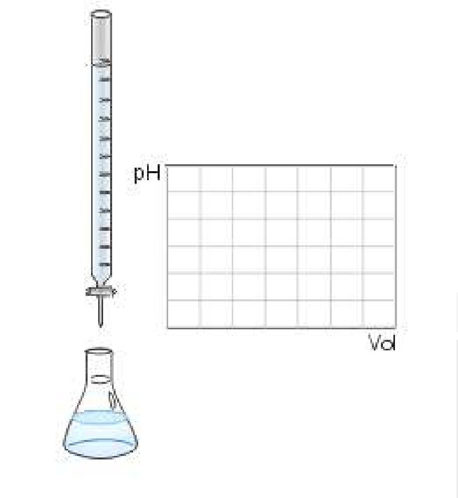
a. Write a balanced equation for the reaction of carbonic acid (H_2CO_3) and potassium hydroxide.

b. Write the net ionic equation for this reaction.

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?

Practice Problems

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

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

Sample Exercise

- How many grams of $Ca(OH)_2$ are needed to neutralize 25.0mL of 0.100M HNO_3?

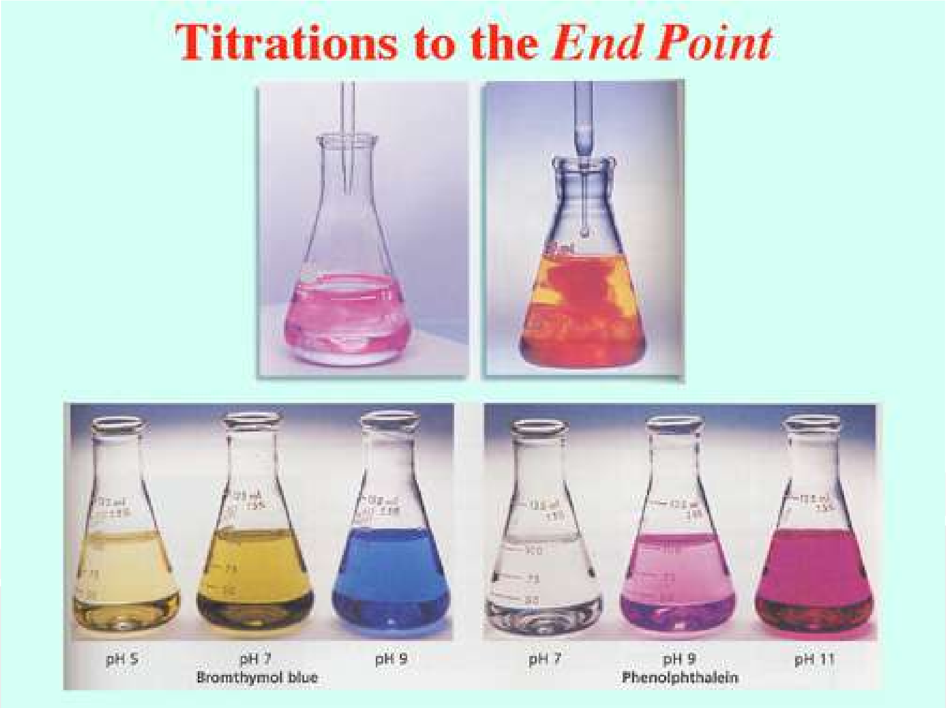
Practice Exercise

a. How many grams of NaOH are needed to neutralize 20.0mL of 0.150M H_2SO_4 solution?

b. How many liters of 0.500M HCl are needed to react completely with 0.100 mol of $Pb(NO_3)_2$, forming a precipitate of $PbCl_2$?

Titration

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



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?

Practice Problems

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

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

Sample Exercise

a. How many grams of chloride ion are in a sample of the water if 20.2mL of 0.100M Ag^+ is needed to react with all the chloride in the sample?

$$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$$

b. If the sample has a mass of 10.0g, what percent Cl^- does it contain?

Practice Exercise

a. A solution that contains Fe^{2+} is titrated with 47.20mL of 0.02240M MnO_4^-. How many moles of MnO_4^- were added to the solution?

$$MnO_4^-{}_{(aq)} + 5Fe^{2+}{}_{(aq)} + 8H^+{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + 5Fe^{3+}{}_{(aq)} + 4H_2O_{(l)}$$

b. How many moles of Fe^{2+} were in the sample?

Practice Exercise

c. How many grams of iron were in the sample?

d. If the sample had a mass of 0.8890g, what is the percentage of iron in the sample?

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

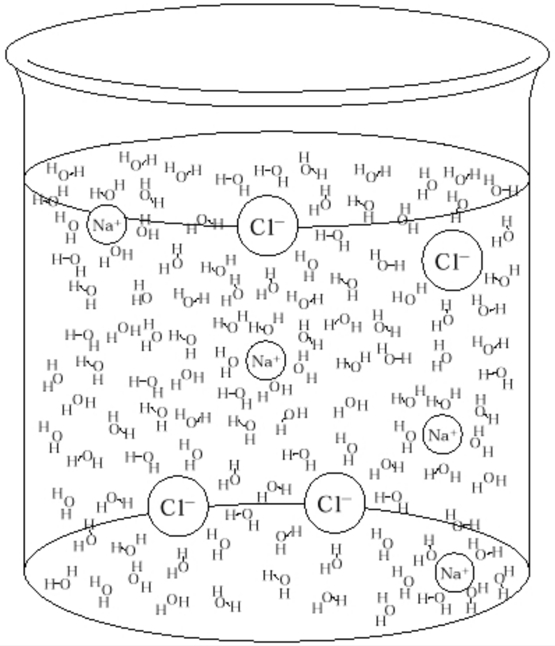
   c. $HNO_3 + Mg(OH)_2 \rightarrow$

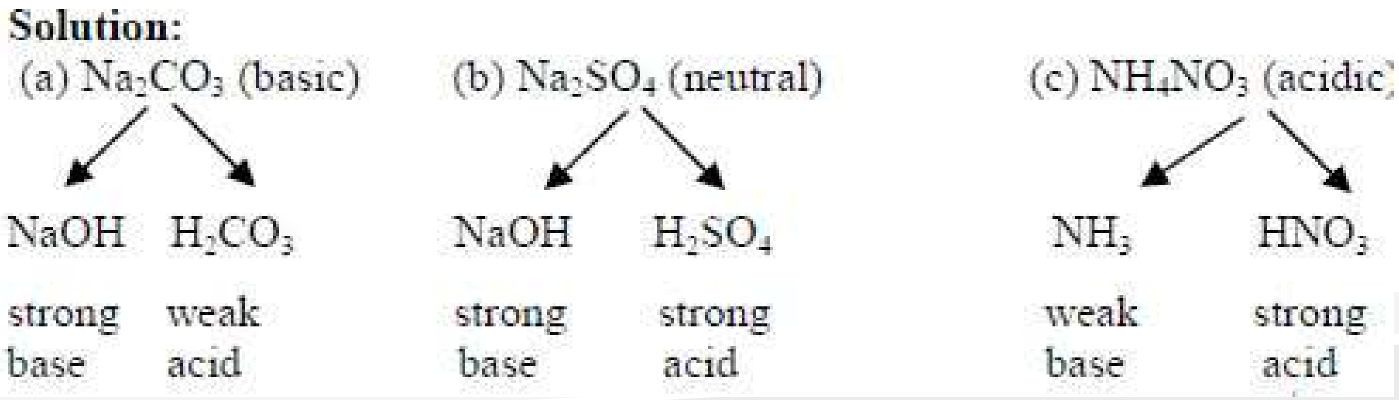
Section 19.5 – Salts in Solutions

- 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

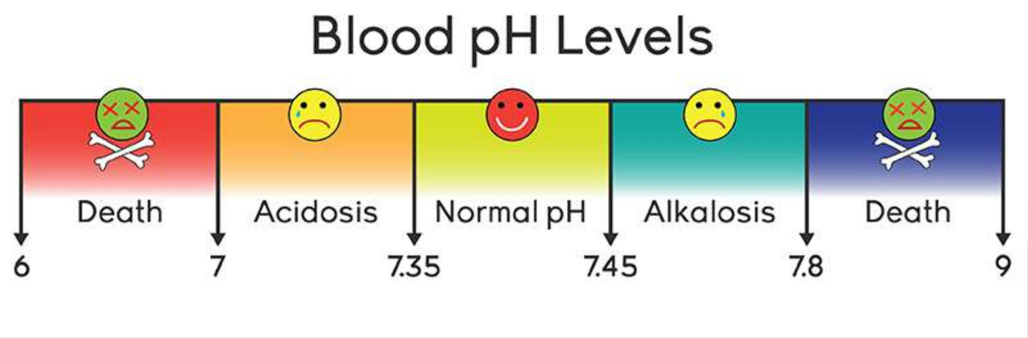
- Strong Acid + Strong Base = Neutral Solution
- Strong Acid + Weak Base = Acidic Solution
- Weak Acid + Strong Base = Basic Solution

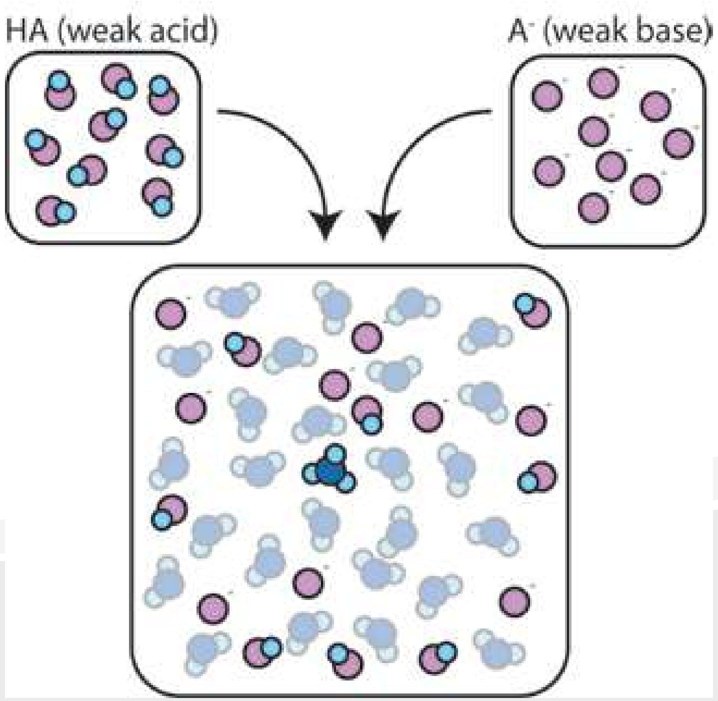




Buffers

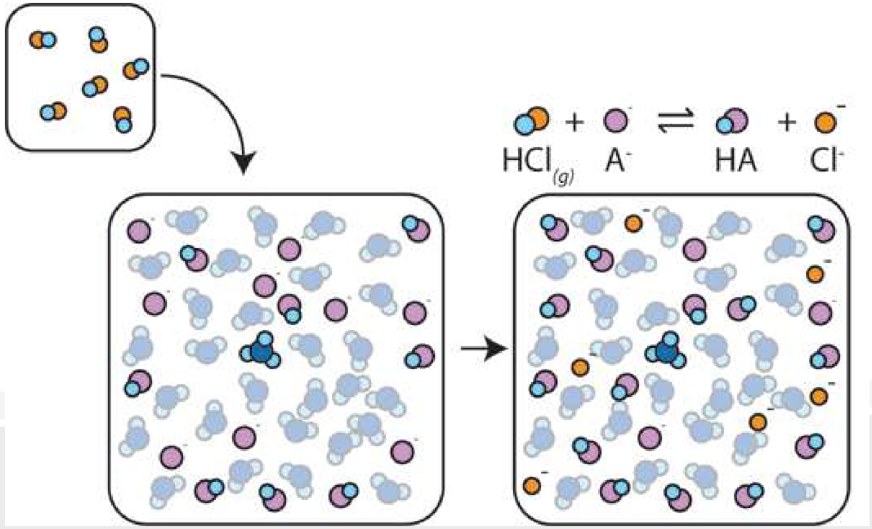
- A buffer is a solution in which the pH remains relatively constant when small amounts of acid or base are added.
- A buffer is made from a weak acid and its conjugate base or a weak base and its conjugate acid.
- A common buffer that you have is your blood.

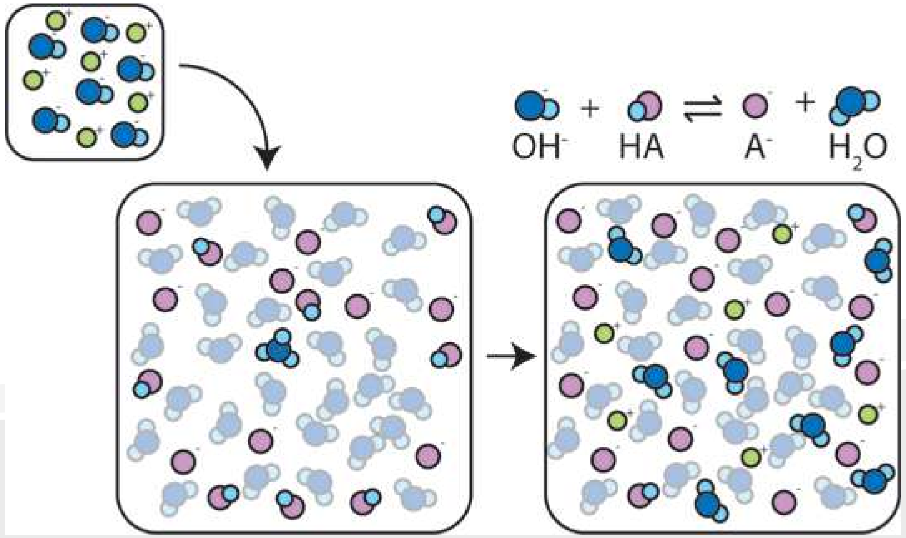




Buffer

- 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. $(NH_4)_2SO_4$