Solutions Basics

16.2 Concentrations of Solutions

Molarity

- **Concentration** of a solution is a measure of the amount of solute that is dissolved in a given quantity of solvent.
  - Solute substance being dissolved
  - Solvent substance doing the dissolving
  - Lemonade solution, solute is lemon juice and sugar, solvent is water.
- **Dilute solution** is on that contains a small amount of solute
- **Concentrated solution** contains a large amount of solute
  - 1 g NaCl per 100 g H2O would be dilute when compared to 30 g NaCl per 100 g H2O (a concentrated solution)

Molarity

- **Molarity (M)** is the number of moles of solute dissolved in one liter of solution
- **To calculate the molarity of a solution, divide the moles of solute by the volume of the solution.**

$$\textbf{Molarity (M)} = \frac{\textbf{moles of solute}}{\textbf{liters of solution}}$$

- Molarity is also known as molar concentration

Molarity Example

- A saline solution contains 0.95 g NaCl in exactly 100 mL of solution. What is the molarity of the solution?

$$\textbf{Molarity (M)} = \frac{\textbf{moles of solute}}{\textbf{liters of solution}}$$

Need moles of solute, so turn grams to moles

$$0.95 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.016 \text{ mol NaCl}$$

Need liters of solution

$$100 \text{ ml} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.100 \text{ L}$$

$$\text{Molarity} = \frac{0.016 \text{ mol NaCl}}{0.100 \text{ L}} = 0.16 \text{ M NaCl}$$

Molarity Example

- A solution has a volume of 2.50 L and contains 159g NaCl. What is the molarity of the solution?

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Need moles of solute, so turn grams to moles

$$0159 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 2.72 \text{ mol NaCl}$$

$$\text{Molarity} = \frac{2.72 \text{ mol NaCl}}{2.50 \text{ L}} = 1.09 \text{ M NaCl}$$

Molarity Example

- What mass of NaCl is needed to make a 2.25 M NaCl solution with a volume of 500 mL?

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Need liters of solution

$$500 \text{ ml} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.500 \text{ L}$$

Molarity x Liters = moles

2.25 M NaCl x 0.500 L = 1.13 mol NaCl

Turn mole to grams

$$1.13 \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol L}} = 66.0 \text{ g NaCl}$$

Molarity Practice (Chemical Formulas)

1. NaCl
2. Na_2S
3. HI
4. $Al(OH)_3$
5. H_2SO_4
6. $Mg_3(PO_4)_2$
7. NaCl
8. CaS

Molarity Practice (Final ANSWERS)

1. 0.78 M NaCl
2. 0.12 M Na_2S
3. 153g HI
4. 0.189 mol $Al(OH)_3$
5. 5.0 L of H_2SO_4 solution
6. 3260 g $Mg_3(PO_4)_2$
7. 13.2 M NaCl
8. 2660 mL of CaS solution

Acids, Bases, and Salts

Chapter 19

19.1 Acid- Base Theories

- Properties of Acids
  - Taste sour
  - React with metals to form H_2 gas
  - Will change the color of and acid-base indicator
  - Can be strong or weak electrolytes in aqueous solutions
- Properties of Bases
  - Taste bitter
  - feel slippery
  - Will change the color of and acid-base indicator
  - Can be strong or weak electrolytes in aqueous solutions

Examples

- Acid: citric acid

- Base: milk of magnesia

Magnesium hydroxide

Lewis Acids and Bases

- Gilbert Lewis's theory of acids and bases was an extension of his concept of electron pairs
- This is the broadest acid/base definition used.
- Lewis acids accepts a pair of electrons
- Lewis bases donate a pair of electrons

Arrhenius Acids and Bases

- Svante Arrhenius in1887 posed a way of defining acids and bases
- <u>Arrhenius Acids</u>
  - Are hydrogen-containing compounds that ionize to yield hydrogen ions (H+)

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

$$NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$$

Types of acids

- Monoprotic acids are acids that contain one ionizable hydrogen.

  Nitric acid: HNO_3

- Diprotice acids are acids that contain two ionizable hydrogens.

  Sulfuric acid: H_2SO_4

- Triprotic acids are acids that contain three ionizable hydrogens.

  Phosphoric acid: H_3PO_4

- Arrhenius Bases
  - Are compounds that ionize to yield hydroxide ions (OH^-) in aqueous solutions

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

$$CH_3COO^- + H_2O \rightarrow CH_3COOH + OH^-$$

Brønsted-Lowry Acids and Bases

- In 1923 Johannes Brønsted and Thomas Lowry proposed a new definition
- Is a more comprehensive definition of acids and bases
- <u>Brønsted-Lowry Acids</u>
  - Is a hydrogen-ion donor
- <u>Brønsted-Lowry Bases</u>
  - Is a hydrogen-ion acceptor

Acid-Base Definitions

Type	Acid	Base
Arrhenius	H^+ producer	OH^- producer
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	Electron-pair acceptor	Electron pair donor

Conjugate Acids and Bases

- Conjugate acid is the particle formed when a base gains a hydrogen ion.
- Conjugate base is the particle formed when an acid has donated a hydrogen ion.
- Conjugate acid-base pair consists of two substance related by the loss or gain of a single hydrogen ion.
- When a water molecule gains a hydrogen ion it becomes a positively charged hydronium ion (H_3O^+).
- When water molecule looses a hydrognet ion it becomes the negativly charged hydroxide ion (OH^{-1})

Conjugate Acid and Bases

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

Acid + Base → Conjugate Acid + Conjugate Base

$$NH_3 + H_2O \rightarrow NH^{+4} + OH^-$$

Base + Acid → Conjugate Acid + Conjugate Base

- A substance that can act as both an acid and a base is said to be amphoteric. (example: Water)

Conjugate acids and base

- What are the conjugate bases of these acids?

a) $HClO_4$ b) HSO_4^- c) $HC_2H_3O_2$ d) H_2S

a) ClO_4^- b) SO_4^{2-} c) $C_2H_3O_2^-$ d) HS^-

- What are the conjugate acids of these bases

a) SO_4^{2-} b) NH_3 c) F^- d) NO_3^-

a) HSO_4^- b) NH_4^+ c) HF d) HNO_3

19.1 Review

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

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

The pH Concept

- The **pH** of a solution is the negative logarithm of the hydrogen-ion concentration.
- $pH = -\log [H^+]$
- For example a neutral solution has $[H^+]$ of 1.0×10^{-7} so the pH is calculated

$$pH = -\log (1.0 \times 10^{-7}) = 7.00$$

- The **pOH** of a solution is the negative logarithm of the hydroxide concentration.
- $pOH = -\log [OH^{-1}]$

pH and Significant Figures

- A $[H^+]$ of 6.0×10^{-5} has two significant figures
- The pH is recorded with two decimal places 4.22

- A $[H^+]$ of $6. \times 10^{-5}$ has one significant figures
- The pH is recorded with one decimal places 4.2

Calculating pH practice

1. $[H^+]$ of 1.0×10^{-11}

   $pH = -\log(1.0 \times 10^{-11}) = 11.00$

2. $[H^+]$ of 6.0×10^{-5}

   $pH = -\log(6.0 \times 10^{-5}) = 4.22$

3. $[H^+]$ of 4.0×10^{-3}

   $pH = -\log(4.0 \times 10^{-3}) = 2.40$

4. $[H^+]$ of 9.0×10^{-9}

   $pH = -\log(9.0 \times 10^{-9}) = 8.05$

Calculating pOH

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

1. $[OH^-]$ of 1.0×10^{-3}

   $pOH = -\log (1.0 \times 10^{-3}) = 3.00$

2. $[OH^-]$ of 6.0×10^{-7}

   $pOH = -\log (6.0 \times 10^{-7}) = 6.22$

3. $[OH^-]$ of 2.0×10^{-4}

   $pOH = -\log (2.0 \times 10^{-4}) = 3.70$

4. $[OH^-]$ of 7.2×10^{-9}

   $pOH = -\log (7.2 \times 10^{-9}) = .14$

- A solution in which $[H^+]$ is **greater** than 1.0×10^{-7} has a pH less than 7.0 and is **acidic**.
- The pH of a neutral solution is 7.0
- A solution in which $[H^+]$ is **less** than 1.0×10^{-7} has a pH greater than 7.0 and is **basic.**
- A simple relationship between pH and pOH allows you to calculate the other if one is known.

$$\mathbf{pH + pOH = 14}$$

Calculating concentration from pH/pOH

- Using pH to find $[H^+]$ since pH= -log $[H^+]$

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

- Calculate the [H+] for the solution

pH= 3.00

$[H^+]= 10^{(-pH)}$ $[H^+] = 10^{(-3.00)} = 1.0 \times 10^{-3}$ M H^+

0.00100 M H^+

- Using pH to find $[OH^{-1}]$ since pOH= -log $[OH^{-1}]$

  Then $[OH^{-1}] = 10^{(-pOH)}$

- Calculate the $[OH^{-1}]$ for the solution

pOH= 7.67

$[OH^{-1}]= 10^{(-pOH)}$ $[OH^{-1}] = 10^{(-7.67)} = 2.14 \times 10^{-8}$

pH to pOH

- If $[H^{+1}]$ is 2.5×10^{-4} what is the pOH

  $pH = -\log [2.5 \times 10^{-4}] = 3.60$

  $pOH = 14 - 3.60 = 10.40$

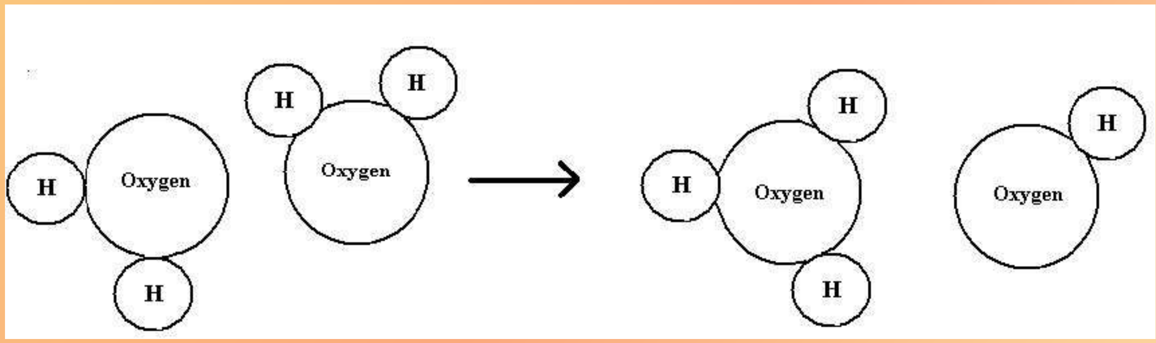
- If the [OH-1] is 6.8×10^{-12} what is the pH

  $pOH = -\log [6.8 \times 10^{-12}] = 11.17$

  $pH = 14 - 11.17 = 2.83$

19.2 Hydrogen Ions and Acidity

- The reaction in which water molecules produce ions is called the self-ionization of water
- $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$
- $2H_2O(l) \rightleftharpoons H^3O^+(aq) + OH^-(aq)$



- Self-ionization of water occurs to a very small extent, in pure water at 25°C.
- Any aqueous solution in which the $[H^+]$ and $[OH^-]$ are equal is called a neutral solution.
- In aqueous solution **$[H^+] \times [OH^-] = 1.0 \times 10^{-14}$**
  - This is called the ion-product constant for water (K_w).
  - The concentrations may change but the product always equals 1.0×10^{-14} for water.

Acidic Solutions

- An acidic solution is one in which the $[H^+]$ is greater than the $[OH^-]$.
- The $[H^+]$ of an acidic solution is greater than 1.0×10^{-7} M

$$HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

Basic Solution

- An basic solution is one in which the $[H^+]$ is less than the $[OH^-]$.
- The $[H^+]$ of an acidic solution is less than
  1.0×10^{-7} M

$$NaOH \xrightarrow{H_2O} Na^+ + OH^-$$

- Basic solutions are also known as alkaline solutions.

- What is the $[OH^-]$ if the $[H^+]$ is 1.0×10^{-3}?

remember: $[OH^-] \times [H^+] = 1.0 \times 10^{-14}$

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

- What is the $[H^+]$ if the $[OH^-]$ is 1.0×10^{-8} ?

$$[H^+] = 1.0 \times 10^{-14} / 1.0 \times 10^{-8} = 1.0 \times 10^{-6}$$

- Classify each solution as acidic, basic, or neutral:

a) $[H^+] = 6.0 \times 10^{-10}$

b) $[OH^-] = 3.0 \times 10^{-4}$

c) $[H^+] = 6.0 \times 10^{-7}$

d) $[H^+] = 1.0 \times 10^{-7}$

- What is the $[H^+]$ of a solution with a pOH of 3.12? Is it acidic, basic, or neutral?

pH= 14- pOH   pH= 14 – 3.12 = 10.88

$[H^+]= 10^{(-pH)}$ $[H^+] = 10^{(-10.88)} = 1.3 \times 10^{-11}$

it is a Basic solution

- What is the $[H^+]$ of a solution with a pOH of 9.18? Is it acidic, basic, or neutral?

pH= 14- pOH   pH= 14 – 9.18 = 4.82

$[H^+]= 10^{(-pH)}$ $[H^+] = 10^{(-4.82)} = 1.5 \times 10^{-5}$

it is an acidic solution

Concentration to pH and back

- Possible equations

$pH = -\log [H^{+1}]$ $pOH = -\log [OH^{-1}]$ $pH + pOH = 14$

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

- If the $[OH^{-1}] = 4.68 \times 10^{-3}$ determine the pOH, pH, $[H^{+1}]$ and if acidic/basic/neutral.

- pOH $= -\log [4.68 \times 10^{-3}] = 2.33$
- *pH* $= 14 - 2.33 = 11.67$
- $[H^{+1}] = 10^{-11.67} = 2.14 \times 10^{-12}$
- Substance is Basic because pH is 11.67

19.3 Strengths of Acids and Bases

- **Strong acids** are completely ionized in aqueous solutions.

$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$ *100% ionized*

- **Weak acids** only slightly ionize in aqueous solutions.

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

Less than 0.4% ionized

Six strong acids!!

- HCl hydrochloric acid
- HBr hydrobromic acid
- HI hydroiodic acid
- HNO_3 nitric acid
- H_2SO_4 sulfuric acid
- $HClO_4$ perchloric acid

All other acids are considered weak acids!!

Acid Dissociation Constant

- An **acid dissociation constant (K_a)** is the ratio of the concentration of the dissociated form of an acid to the concentration of the undissociated form.

$$K_a = \frac{[H_3O^+] \text{ x } [\text{dissocated acid}]}{[\text{acid}]} = \frac{[H^+][A^-]}{[HA]}$$

- Weak acids have small K_a values.
- The stronger an acid is the larger is its K_a value.

Base Dissociation Constant

- **Strong bases** dissociate completely into metal ions and hydroxide ions in aqueous solution.

- **Weak bases** react with water to form the hydroxide ion and the conjugate acid of the base.

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

Eight strong bases

- LiOH - lithium hydroxide
- NaOH - sodium hydroxide
- KOH - potassium hydroxide
- RbOH - rubidium hydroxide
- CsOH - cesium hydroxide
- $Ca(OH)_2$ - calcium hydroxide
- $Sr(OH)_2$ - strontium hydroxide
- $Ba(OH)_2$ - barium hydroxide

Base Dissociation Constant

- The **base dissociation constant (K_b)** is the ratio of the concentrations of the conjugate acid time the concentration of the hydroxide ion to the concentration of the base.

$$K_b = \frac{[\text{conjugate acid}] \times [OH^-]}{[\text{base}]}$$

Calculation a Dissociation Constant

Only for weak acids

- At equilibrium $[H^+]=[A^-]$ and $[HA] = M - [H^+]$
- M is the molarity of the solution

- A 0.1500 M solution of acid is only partially ionized. The $[H^+]$ is determined to be 2.56×10^{-3} . What is the K_a of this acid solution?

$[H^+]=[A^-]= 2.56 \times 10^{-3}$

$[HA] = 0.1500M - 2.56 \times 10^{-3} =0.1474$

$$K_a= \frac{(2.56 \times 10^{-3})\times(2.56 \times 10^{-3})}{0.1474} = 4.446 \times 10^{-5}$$

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

1. Calculate the K_a of a 0.250 M weak acid solution if the $[H^+]$ equals 9.45×10^{-4}.

$[H^+]=[A^-]= 9.45 \times 10^{-4}$

$[HA]= 0.2500 - 9.45 \times 10^{-4} = 0.249$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[9.45 \times 10^{-4}]\,[9.45 \times 10^{-4}]}{0.249} = 3.59\times10^{-6}$$

Strong Acids or Bases

- $[H^+]$ for strong acids equals molarity of acid

$$[H^+] = \text{M of acid}$$

- $[OH^-]$ for strong bases equals molarity of base

$$[OH-] = \text{M of base}$$

- What is the $[OH^-]$ and $[H^+]$ for the following solutions

a. 0.275 M HCl
b. 0.500 M NaOH
c. 0.200 M HNO_3
d. 0.375 M $Ba(OH)_2$

Remember $[H^+] \times [OH^-] = 1.0 \times 10^{-14}$

a. 0.275 M HCl

$[H^+]= 0.275$ $[OH^-] = 3.64 \times 10^{-14}$

b. 0.500 M NaOH

$[OH^-] =0.500$ $[H^+]= 2.00 \times 10^{-14}$

c. 0.200 M HNO_3

$[H^+]=0.200$ $[OH^-] = 5.00 \times 10^{-14}$

d. 0.375 M $Ba(OH)_2$

$[OH^-] =0.375$ $[H^+]= 2.67 \times 10^{-14}$

19.4 Neutralization Reactions

- When you mix a strong acid with a strong base a neutral solution results.

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

$$2\,HBr_{(aq)} + Ba(OH)_{2\,(aq)} \rightarrow BaBr_{2(aq)} + 2\,H_2O_{(l)}$$

- A <u>neutralization reaction</u> is where an acid and a base react to form a salt and water

Titration

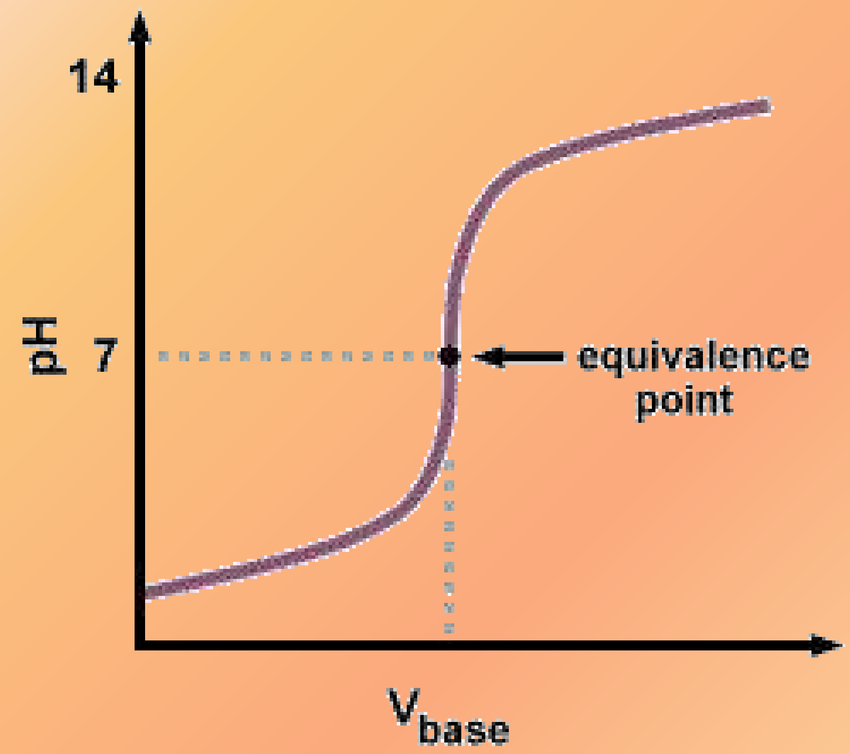
- Acids and bases do not always react in a 1:1 ratio
- The <u>equivalence point</u> is when the number of moles of hydrogen ions equals the number of moles of hydroxide ions.
- How many moles of hydrochloric acid are required to neutralized 0.50 moles of barium hydroxide ?

$$2HCl_{(aq)} + Ba(OH)_{2\,(aq)} \rightarrow BaCl_{2\,(aq)} + 2H_2O_{(l)}$$

$$0.50 \text{ mol } Ba(OH)_{2\,(aq)} \times \frac{2 \text{ mol HCl}}{1 \text{ mol } Ba(OH)_2} = 1.0 \text{ mol HCl}$$

- The process of adding a known amount of a solution of known concentration to determine the concentration of another solution is called titration.
- The solution of known concentration is called the standard solution.
- We use a buret to add the standard solution.
- The solution is added until the indicator changes colors.
- The point at which the indicator changes colors is the end point of the titration.

Titration of a strong acid with a strong base



Titration Calculations – 1st type of problem

- Determine the volume of 2.1 M sodium hydroxide needed to titrate 435 mL of 1.75 M sulfuric acid.

1st : write a balanced chemical equation

2nd : determine the number of moles of reactant (sulfuric acid) available

3rd: use stoichiometry to calculate the number of moles of reactant (sodium hydroxide) needed

4th : use molarity to determine the number of mL needed to complete the neutralization.

- Determine the volume of 2.1 M sodium hydroxide needed to titrate 435 mL of 1.75 M sulfuric acid.

1st : $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

2nd : 0.435L x [1.75 mol/1L]= 0.76125 mol H_2SO_4

3rd: 0.76125mol H_2SO_4 x $\dfrac{\text{2mol NaOH}}{\text{1 mol } H_2SO_4}$ = 1.5225 mol NaOH

4th : 1.5225 mol NaOH x $\dfrac{\text{1 L}}{\text{2.1 mol}}$ = 0.725 L = 725 ml

- Determine the volume of 3.3 M hydrochloric acid needed to titrate 355 mL of 2.50 M magnesium hydroxide .

1st : $2HCl + Mg(OH)_2 \rightarrow MgCl_2 + 2H_2O$

2nd : 0.355L x [2.50 mol/1L]= 0.8875 mol $Mg(OH)_2$

3rd: 0.8875mol $Mg(OH)_2$ x $\frac{2\text{mol HCl}}{1\text{ mol } Mg(OH)_2}$ = 1.775 mol HCL

4th : 1.775mol HCl x $\frac{1\text{ L}}{3.3\text{ mol}}$ = 0.538 L = 538 ml HCl

Titration Calculations – 2nd type of problem

- Titration reveals that 22.5 ml of 2.0 M nitric acid are required to neutralize 20 ml of calcium hydroxide. What is the molarity of the calcium hydroxide?

1st : $2HNO_3 + Ca(OH)_2 \rightarrow Ca(NO_3)_2 + 2H_2O$

2nd : $0.0225\text{L} \times [2.0\ \text{mol}/1\text{L}] = 0.045\ \text{mol}\ HNO_3$

3rd: $$0.045\text{mol}\ HNO_3 \times \frac{1\text{mol}\ Ca(OH)_2}{2\ \text{mol}\ HNO_3} = 0.0225\ \text{mol}\ Ca(OH)_2$$

4th : $$\frac{0.0225\text{mol}\ Ca(OH)_2}{0.020\ \text{L}} = 1.125\ \text{M}\ Ca(OH)_2$$

- Titration reveals that 11.2 ml of 3.5 M barium hydroxide are required to neutralize 50 ml of perchloric acid. What is the molarity of the perchloric acid?

1st : $2HClO_4 + Ba(OH)_2 \rightarrow Ba(ClO_4)_2 + 2H_2O$

2nd : $0.0112 L \times [3.5 \text{ mol}/1 L] = 0.0392 \text{ mol } Ba(OH)_2$

3rd: $0.0392 \text{mol } Ba(OH)_2 \times \frac{2 \text{ mol } HClO_4}{1 \text{ mol } Ba(OH)_2} = 0.0784 \text{ mol } HClO_4$

4th : $\frac{0.0784 \text{ mol } HClO_4}{0.050 L} = 1.56 \text{ M } HClO_4$

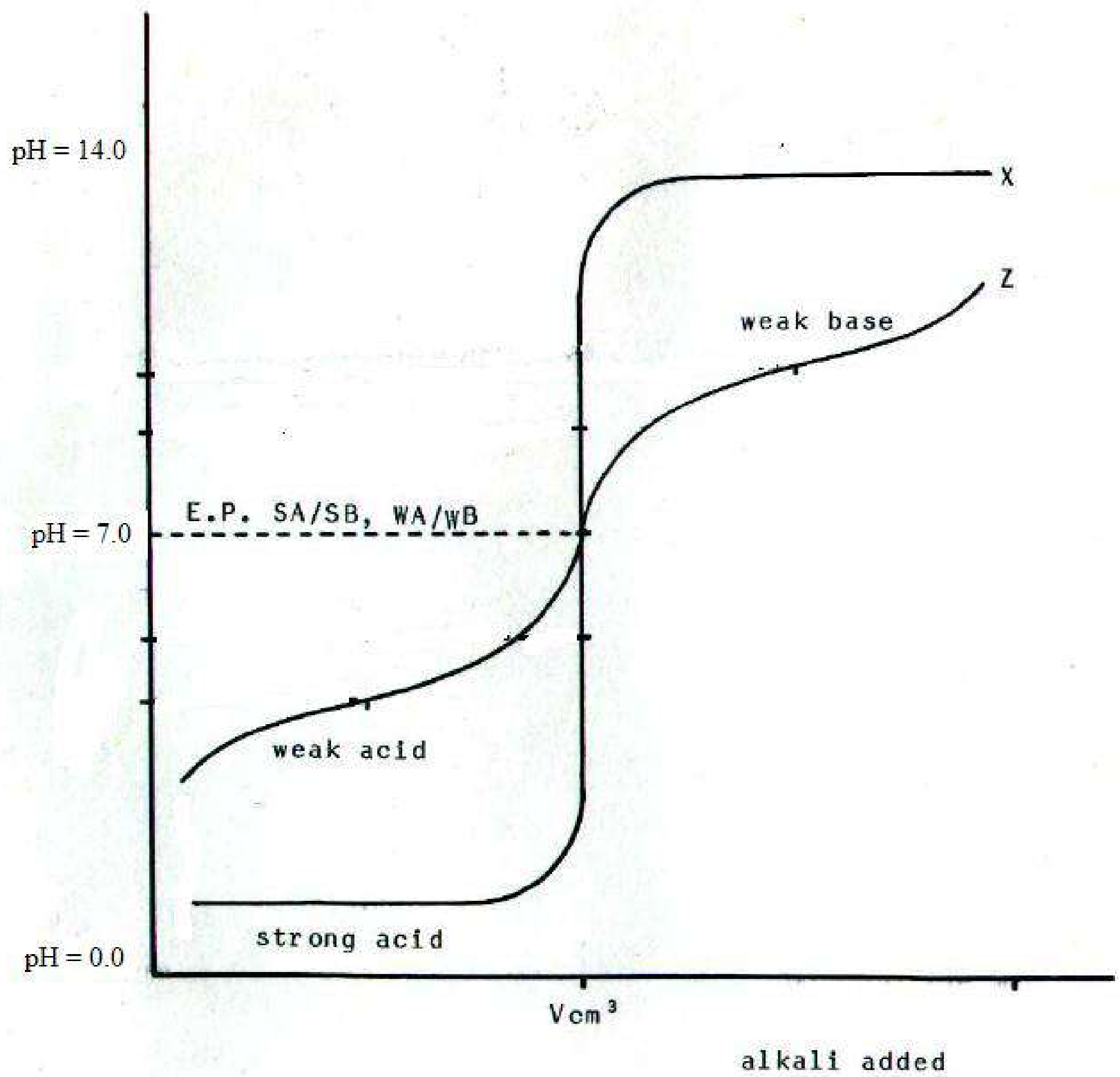
19.5

- Salt hydrolysis the cations or anions of a dissociated salt remove hydrogen ions from or donate hydrogen ions to the water.
- In general, salts that produce acidic solutions contain positive ions that release protons to water
- Salts that produce basic solutions contain negative ions that attract protons from water.

19.5 Salts in Solution

- A salt consists of an anion from an acid and a cation from a base. It forms as the result of a neutralization reaction.
- A <u>buffer</u> is a solution in which the pH remains relatively constant when small amount of acid or base are added.
- A buffer is a solution of a weak acid and one of its salts, or a solution of a weak base and one of its salts.
- Buffers are able to resist drastic changes in pH.

- The buffering capacity is the amount of acid or base that can be added to a buffer solution before a significant change in pH occurs.
- Buffer systems are crucial in maintaining human blood pH within a narrow range.


pH = 14.0
X
Z
weak base
E.P. SA/SB, WA/WB
pH = 7.0
weak acid
strong acid
pH = 0.0
Vcm³
alkali added