Analytical Chemistry
for Pharmacy Students

By

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Lecture 5 - 6

QUANTITATIVE ANALYTICAL CHEMISTRY

- What is Analytical Chemistry
- Defining the problem
- Some of the common measuring techniques involves
- 1 - Acid-Base titration (neutralization reaction)
- 2 - Reduction-Oxidation titration (redox reaction)
- 3 - Precipitation titration
- 4 - Complexometric titration
- 5 - Gravimetric titration
- 6 - Instrumental analysis involving
-   a - spectroscopic methods
-   b - electrochemical methods
-   c - chromatographic methods

ACID-BASE EQUILIBRIA

- **Chemical reactions : The rate concept**
- **The molar equilibrium constant**
  (for dilute solution)
- **the rate of the forward reaction**
  $$\text{rate f} = k_f [A]^a [B]^b$$
- **the backward reaction has also a rate equal to**
  $$\text{rate b} = k_b [C]^c [D]^d$$

- $$K = \frac{k_f}{k_b} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

AQUEOUS ACID - BASE THEORIES

- **<u>Arhenius Theory</u>**

  An acid is any substance that ionizes partially or completely in water to give hydrogen ions (protons) .

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

- **<u>Bronsted Theory</u>**

  An acid is any substance that can donate a proton

  A base is any substance that can accept a proton

$$Acid \quad \leftrightarrow \quad H^+ \quad + \quad Base$$
$$HB \quad \leftrightarrow \quad H^+ \quad + \quad B$$

- **<u>Lewis Theory</u>**

  Lewis acid is substance that have unfilled electron shells , it reacts by accepting an electron pair from a base ,

  e.g. BF_3 , M (cation in complex) , $AlCl_3$

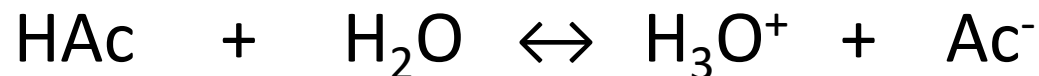# ACID BASE EQUILIBRIUM IN WATER

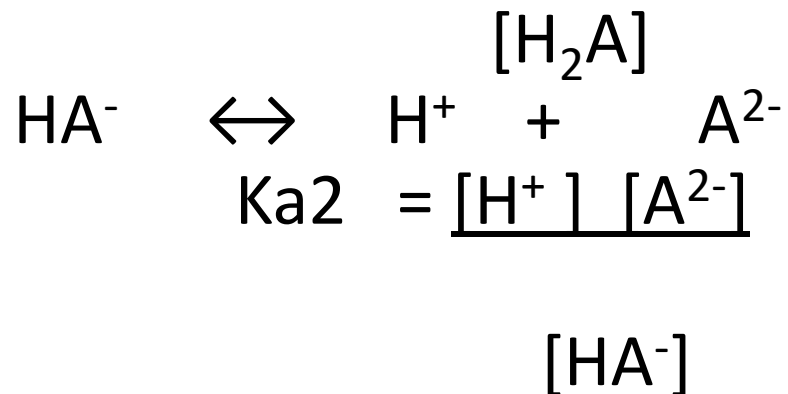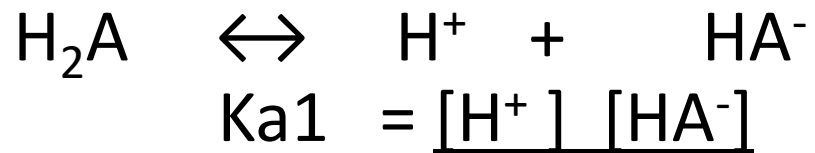| Strong acid	Weak acid	Strong base	Weak base
HCl	CH_3COOH	NaOH	NH_3
HNO_3	C_6H_5OH	$NaC_2H_3O_2$	$C_6H_5NH_2$
H_2SO_4	HCOOH		
$HClO_4$			

- Hydrochloric acid is a strong acid , it is completely ionized

$$HCl \rightarrow H^+ + Cl^-$$

- Acetic acid is a weak acid , it is partially ionized

$$HAc + H_2O \leftrightarrow H_3O^+ + Ac^-$$

- When a polyprotic acid is dissolved in water , the various hydrogen atoms undergo ionization to different extent .
- For example ,
- for the dibasic acid H_2A the primary and secondary ionization can be represented by the equations

$$H_2A \longleftrightarrow H^+ + HA^-$$

$$Ka1 = \frac{[H^+]\,[HA^-]}{[H_2A]}$$

$$HA^- \longleftrightarrow H^+ + A^{2-}$$

$$Ka2 = \frac{[H^+]\,[A^{2-}]}{[HA^-]}$$

- The greater the value ka1 relative to ka2 , the smaller will be the secondary dissociation

- **Example for polyprotic acid is phosphoric acid**

- $H_3PO_4 \rightarrow H_2PO_4^- + H^+$ ka1 $= 1.1 \times 10^{-2}$

  $H_2PO_4^- \rightarrow HPO_4^{2-} + H^+$ ka2 $= 2.0 \times 10^{-7}$

  $HPO_4^{2-} \rightarrow PO_4^{3-} + H^+$ ka3 $= 3.6 \times 10^{-13}$

The Dissociation Constant of Water

- $H_2O \quad \longleftrightarrow \quad H^+ + OH^-$

  $$Kw = \frac{[H^+]\ [OH^-]}{[H_2O]}$$

- The ionic product of water

  $Kw = [H^+]\ [OH^-] = 1.0 \times 10^{-14}$ at 25 °C.

- in pure water

  $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ (at 25 °C).

- A solution in which $[H^+] = [OH^-]$ is said to be neutral ,

- a solution is said to be acidic when $[H^+] > [OH^-]$,

- it is basic when the reverse is the case .

The pH Scale :

- pH   = - log [H^+]
- pOH = -log [OH^-]
- Kw   = [H^+] [OH^-] = 1.0 x 10^{-14}
- pKw =pH + pOH   = 14
- In neutral solution pH = pOH = 7 .
- Acidic solutions have pH values less than 7 , while alkaline solutions have pH values greater than 7 ,

- **For any acid base pair**

  Ka x Kb = Kw

- **pKa + pKb   = 14**

pH of Acids and Bases

1 - pH of solutions strong acids and bases :

- Example : calculate the pH of 2.0×10^{-3} M solution of HCl .

$$[H^+] = 2.0 \times 10^{-3} \text{ M}$$

$$pH = - \log (2.0 \times 10^{-3})$$

$$= 3 - \log 2$$

$$= 3 - 0.3 = 2.7$$

- Example : calculate the pH of 0.01M solution of HCl

$$[H+] = 0.01 = 10^{-2} \, , \, pH = -\log 10^{-2} = 2$$

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

From the equation $\quad$ Kw $\quad = \ [H^+] [OH^-]$

$\quad\quad\quad\quad\quad\quad\quad\quad$ pKw $\ = \ $ pH $\ + \ $ pOH

$\quad\quad\quad\quad\quad\quad\quad\quad$ 14 $\quad = \ $ pH $\ + \ $ pOH

$\quad\quad\quad\quad\quad\quad\quad\quad$ pH $\quad = \ $ 14 $\ - \ $ pOH

$\quad\quad\quad\quad\quad\quad\quad\quad$ pOH $\ = \ $ 14 $\ - \ $ pH

- Examples : Calculate pOH and pH of 5.o $\ x \ 10^{-2}$ M solution of NaOH .

$\quad\quad\quad\quad$ $[OH^-] \ = \ 5.0 \ x \ 10^{-2}$ M

$\quad\quad\quad\quad$ pOH $\ = -\log 5.0 \ x \ 10^{-2} \ = \ 2 \ - \log 5$

$\quad\quad\quad\quad\quad\quad\quad\quad$ $= \ 2 \ - \ 0.7 \quad\quad = \ 1.3$

$\quad\quad\quad\quad$ pH $\quad = \ 14 \ - \ 1.3 \quad\quad = \ 12.7$

- **2 - pH of Weak Acids and Weak Bases**

  a- pH of Weak Acids eg.acetic acid

    HAc $\leftrightarrow$ H^+ + Ac^-

 Ka = [H^+] [Ac^-] = 1.75 x 10^{-5}
          [HAc]

  pH = ½ pCa + ½ pKa

Example :
    calculate the pH and pOH of 1.0 x 10^{-3} solution of acetic acid

                    HAc $\leftrightarrow$ H^+ + Ac^-
- initial concentration   1.0 x 10^{-3} 0   0
- change   -x   + x   + x
- equil.conc.   1.0 x 10^{-3} - x   x   x

          Ka   = $\underline{(x).(x)}$ = 1.75 x 10^{-5}
                    1.0 x 10^{-3}

          x^2 / 1.0 x 10^{-3} = 1.75 x 10^{-5}

          x = 1.32 x 10^{-4} M = [H^+]

   pH = - log 1.32 x 10^{-4} = 4 - log 1.32 = 4 - 0.12 = 3.88
   pOH = 14 - 3.88 = 10 .12

B - for the basicity constant Kb ,eg.amm.hydroxide

$$pH = pKw - \tfrac{1}{2} pCb - \tfrac{1}{2} pKb$$
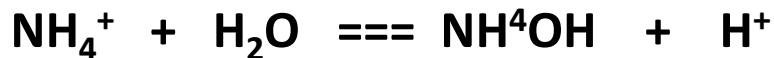
3- pH of Salt solutions

i - Salts of weak acids are those derived from weak acids and strong bases e.g. sodium acetate

$Ac^- + H_2O === HAc + OH^-$
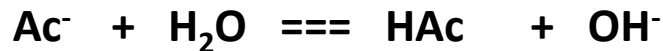
$pH = \tfrac{1}{2} pKw + \tfrac{1}{2} pKa - \tfrac{1}{2} pCs$

ii - Salts of weak bases are those derived from strong acids and weak bases e.g. ammonium chloride ,

$NH_4^+ + H_2O === NH^4OH + H^+$

it behaves as a weak acid and its salts .

$pH = \tfrac{1}{2} pKw - \tfrac{1}{2} pKb - \tfrac{1}{2} pCs$

Buffers

- A buffer is defined as a solution that resists changes in pH when a small amount of an acid or base is added or when the solution is diluted .

- A buffer solution consists of a mixture of weak acid and its conjugate base (HAc and Sodium acetate or a weak base and its conjugate acid (NH_4OH and NH_4Cl) .

The buffer action :

- For buffer system **<u>consisting of weak acid and its salt</u> ;** addition of a strong acid :

$$H+ + Ac- \leftrightarrow HAc \text{ (weak acid)}$$

- The strong acid is , therefor , converted to a weakly dissociated acid and the pH practically does not change . Addition of a strong base :

$$OH^- + HAc \leftrightarrow H_2O + Ac^- \text{ (weak base)}$$

- For buffer system **<u>consisting of a weak base and its salt ;</u>** upon addition of an acid :

$$H^+ + NH_4OH \leftrightarrow NH_4^+ + H_2O$$

- upon addition of a base :

$$OH^- + NH_4Cl \leftrightarrow NH_4 OH + Cl^-$$

Calculation of the pH of a buffer solution ;Henderson equation :

A – Acedic buffer

$$pH = pKa - \log \frac{[HAc]}{[Ac^-]}$$

$$pH = pKa + \log \frac{[\text{conjugate base } [salt]}{[Acid]}$$

Calculate the pH of a solution (buffer) prepared by adding 10 ml of 0.1M acetic acid to 20 ml of 0.1M sodium acetate . (kA = 1.75 x 10^{-5})

- **mmol.acetic acid = 10 ml x 0.1M = 1 mmol.**
- **Final volume = 20 + 10 = 30 ml**
- **cHA = 1 mmol./30ml = 0.33 mmol./ml**
- **cAc$^-$ = 20 ml x 0.1 M / 30 ml = 0.067 mmol./ml**
- **pH = pKa + log $\underline{[Ac^-]}$**

    [HAc]
- **   pH = - log (1.75 x 10^{-5}) + log (0.067/0.33)**

   = 4.76 + log 2 = 5.06

Buffer Capacity :

- the amount of acid or base that can be added without causing a large change in the pH .
- This is determined by the buffer ratio [Salt]/[Acid] .
- The buffer capacity is maximum when [Salt] = [Acid] ,
- that is when pH = pka + log (1/1)

$$pH = pKa$$

B – Basic Buffer "

   system from weak bases and their salts

   e.g. NH_4OH / NH_4Cl

- **similarly the pH of such buffer solution is given by**

   pOH = pKb + log [Salt] / [Base]

   pH   = pKw - pKb - log [Salt] / [Base]

ACID - BASE TITRATIONS

- Many compounds , both inorganic and organic , are either acids or bases , and can be titrated with standard solution of a strong acid or a strong base . The end point can be detected by an indicator or pH - meter .

- ***Titration*** refers to quantitative chemical analysis carried out by measuring the volume of a solution of accurately known concentration which is required to react quantitatively with a measured volume of a solution of the substance to be determined . The solution of accurately known strength or concentration is called ***standard solution*** , the substance in the sample to be determined is called ***analyte*** . The weight of the analyte is calculated from the volume of the standard solution used and the chemical equation and relative molecular mass of the reacting compounds " *Volumetric Analysis* ".
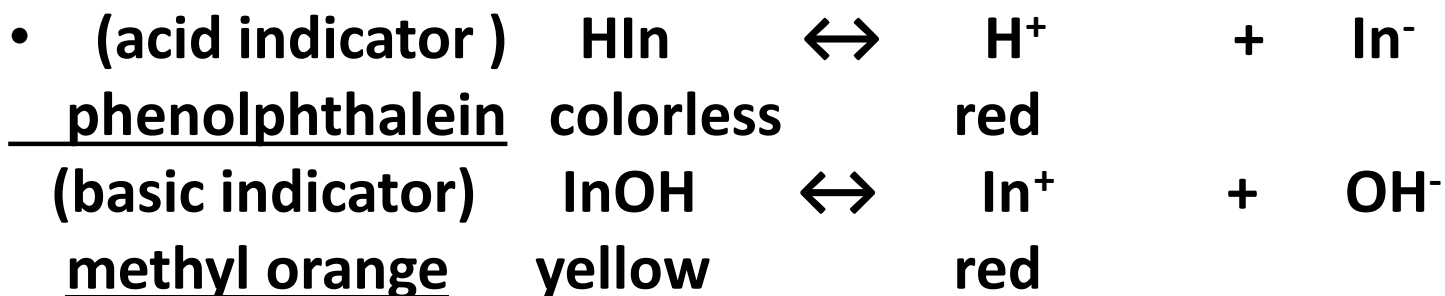
- In titremitric analysis the standard solution is called titrant , the process of adding the standard solution from the burette until the reaction is complete is termed _titration_ , the point at which the reaction is complete is called the **_equivalence point_** or the **_end point_** . The completion of the titration is detected by following up the changes in some physical properties such as pH , potential , conductance ,..etc. , or by the addition of an auxiliary reagent known as _indicator_ .

Detection of the End Point
" Indicators "

- **Types of acid -base indicators :**

a - Color indicators

1 - Universal indicators

2 - mixed indicators

3 - Screened indicators

b - Turbidity indicators

c - Fluorescence indicators

Theory of color indicator

- An indicator for acid base titration is usually weak acid or weak base that is highly colored organic compound . The color of the ionized form is markedly different from that of the unionized form

- (acid indicator)   HIn $\leftrightarrow$ H^+ + In^-

  phenolphthalein   colorless   red

  (basic indicator)   InOH $\leftrightarrow$ In^+ + OH^-

  methyl orange   yellow   red

- For acid indicator

  pH = pKind +   log $[In^-]$ / [HIn]

  pH = pKind +   log [ionized]/[unionized]

  pH = pKind   +   log [basic color]/[acid color]

- **For basic indicator**

  pH = pKw - pK - log [In-]/[InOH]

  pH = pKw - pK - log [ionized]/[unionized]

- **tables for the value of (pKw - pK) for basic indicators usually express this value as pKind.**

  pH = pKind + log [basic color]/[acid color]

- **The indicator changes color over a pH range . <u>The transition range</u> depends on the ability of the observer to detect small color changes . With indicators in which both forms are colored , only one color is observed if the ratio of the concentration of the two forms is 10 : 1 . We can calculate the pH transition range required to go from one color to the other ; if color of the unionized form is seen , then**

     [In-]/[HIn] = 1/10

    pH = pKind + log 1/10

      = pKind - 1

 if the color of the ionized form is observed ,

 then [In-]/[HIn] = 10/1

    pH = pKind + log 10/1

    pH = pKind + 1

- **so most indicators require a transition range from pK -1 to pK + 1**

  i.e.   pH = pKind $\pm$ 1

- *effective range* of the indicator . The interval during which shades of colors are produced extending between the two color forms is called *transition interval* . For a good indicator the effective range its effective range as well as its transition interval should be short , i.e. pH change of 2 . in the midpoint in this transition range the concentration of the two forms are equal i.e. pH = pKind + log 1/1

$$pH = pKind$$

- i.e. the pK of the indicator should be close to the pH of the end point . Basic indicator should be selected that pH = 14 - pKb .

ACID - BASE TITRATION CURVES

1 - Strong Acid - Strong Base Titration

A titration curve is constructed by plotting the pH of the solution as a function of the volume of the titrant added .

Example : titration of HCl against NaOH

- At the beginning The equation used to calculate pH

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

- and during the titration

$$[H+] = \frac{\text{remaining volume of HCl x M}}{\text{total volume}}$$

- at the equivalent point

$$pH = pOH = \frac{1}{2} pKw = 7$$

- after the equivalence point , the excess NaOH added controls the pH of the medium , thus the pH = pKw - pCb

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

$$[OH^-] = \frac{\text{excess volume of NaOH x M}}{\text{total volume}}$$

- **Titration curve for 100 ml 0.1 M HCl titrated with 0.1 M NaOH .**

- **At the beginning , 0.1 M HCl**

  $$pH = -\log 10^{-1} = 1$$

- **at 90% neutralization (90 ml NaOH) , only 10% of H+ remains , neglecting volume change**

  $$[H+] = 10 \text{ ml} \times 0.1 / 100 \text{ ml} = 10^{-2}$$

  $$pH = -\log 10^{-2} = 2$$

- **at the equivalence point , neutralization is complete and neutral solution of NaCl remains   pH = 7**

- **after neutralization , excess OH⁻ rapidly increases from 10^{-7} M at equivalent point to reach 10^{-2} and $M 10^{-1}$ M equivalent to pH 12 and 13 and the solution contains NaOH and NaCl , the pH changes markedly very near the end point .**

2 - <u>Weak Acid - Strong Base Titration</u>

Example : the titration of 100 ml of 0.1M acetic acid with 0.1M NaOH solution .

- 1 - The initial pH of 0.1M acetic acid (pka =4.74) is calculated from the equation

$$pH = ½ pka + ½ Ca$$
$$= 2.37 + 0.5 = 2.88$$

- 2 - when 50 ml of 0.1M alkali has been added , some of the acid is converted to sodium acetate forming a buffer system , as the titration proceeds the pH increase as the ratio [salt] / [acid] change , the pH is calculated from the formula for buffers :

$$pH = pka + log [salt] / [acid]$$
$$= 4.74 + log 50 / 50 = 4.74$$

3 - when 90 ml of alkali has been added then

$$pH = 4.74 + \log 90 / 10 = 5.7$$

4 - At the end point the pH of the solution is given by   the formula for salts :

$$pH = \tfrac{1}{2} pkw + \tfrac{1}{2} pka - \tfrac{1}{2} pCs$$

Cs at the end point - log (100 ml x 0.1 M / 200 ml)

$$= 0.05$$

$$pH = 7 + 2.37 - 0.65 = 8.73$$

5 - Beyond the end point

- The sodium acetate in presence of excess OH^- is negligible because excess alkali suppresses its hydrolysis .So when adding 110 ml of the alkali

  mmol OH- = 10 ml x 0.1M

  $[OH^-]$ = 1 / 210 = 0.00476 M

  pOH = -log 0.00467 = 2.04

  pH   = 14 - 2.04   = 11.96

- The titration curve beyond the equivalence point follows that for titration of strong acid . The slowly rising region before the equivalence point is called the buffer region .

- The curve indicates that methyl orange and methyl red are unsuitable , the end point is at pH 7.8 , the most suitable indicator is phenolphthalein which has transition range in the alkali side (pH 8.3 - 10) or thymolphthalein (pH 9.3 - 10.5) . The break in the titration curve is influenced by the pka value of the acid ;

- The detection of the end point becomes more critical problem for weaker acids ; e.g. 0.1M solution with ka $< 10^{-6}$ could not be titrated accurately .

3 - Strong Acid - Weak Base titration

Example : titration of 100 ml 0.1 M NH_3 ($Kb = 4.74 \times 10^{-5}$) with 0.1M HCl .

1 - At the beginning the pH is calculated from the formula for weak base

pH = 14 - ½ pkb - ½ pCb

= 14 - 2.37 - 0.5 = 11.13

2 - During the titration the pH is calculated using the formula for buffer

e.g. when 90 ml 0.1M HCl is added

pH = 14 - pKb + log [base] /[salt]

= 14 - 4.74 + log 10/90 = 8.31

3 - At the equivalent point the pH is given by the formula for salts

pH = ½ pKw - ½ pkb + ½ pCs

= 7 - 2.37 + 0.6 = 5.28

4 - After the end point the solution represents a strong acid solution and its pH can be calculated in the usual manner .

• The curve shows that the end point is at 5.1 , and it is necessary to use an indicator with pH range on the slightly acidic side such as methyl orange or methyl red (pH 3.6 - 5).

- **4 - Weak Acid - Weak Base**

  The titration of weak acid , like acetic acid , with a weak base , like ammonium hydroxide gives a titration curve the main feature of which is that the change of pH near the end point and during the titration curve is very gradual . The calculation of the titration curve is more difficult than the previous three cases . There is no sudden change in pH at the end point . Therefore , such titrations are of little practical value as no indicator will give a sudden change of the color at this point and the change of the indicator is so gradual .