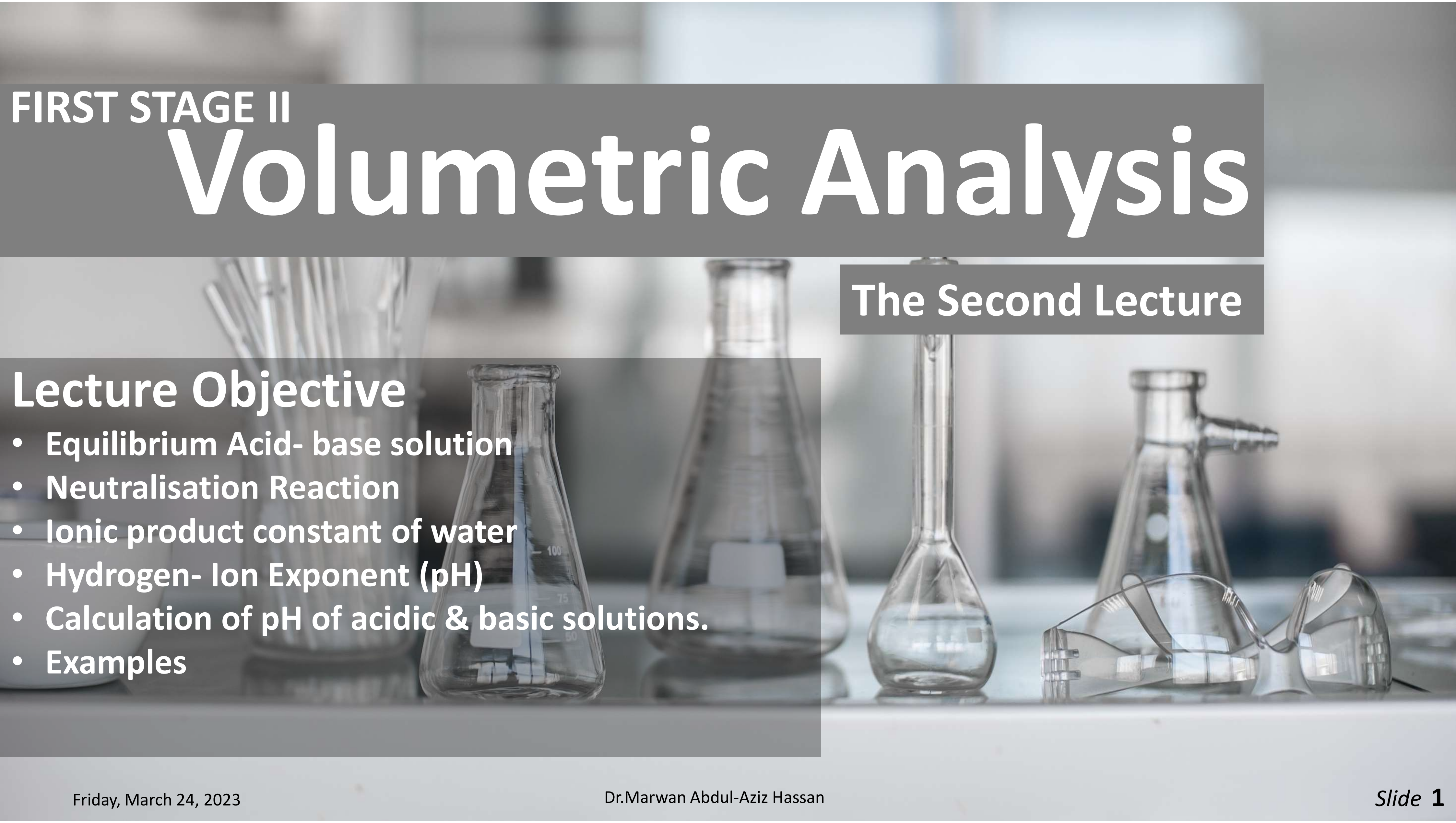

FIRST STAGE II

Volumetric Analysis

The Second Lecture

Lecture Objective

- Equilibrium Acid- base solution
- Neutralisation Reaction
- Ionic product constant of water
- Hydrogen- Ion Exponent (pH)
- Calculation of pH of acidic & basic solutions.
- Examples

Second Lecture

Equilibrium

Neutralisation

K_w

pH

Calculation of pH

Equilibrium

Water (H_2O) – the most important molecule on earth. Even in pure water, there are small amounts of ions from the equilibrium below ("self-ionization of water" or "auto-ionization of water").

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

More accurately:

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

$H_3O^+_{(aq)}$ = hydronium ion; often abbreviated as $H^+_{(aq)}$
$OH^-_{(aq)}$ = hydroxide ion
$[H_3O^+] = [OH^-]$ in pure water

Second Lecture

Equilibrium

Neutralisation

K_w

pH

Calculation of pH

Neutralization

Neutralization: the reaction between an acid and a base called (neutralization reaction) and the water molecular is one of chief products

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

Strengths of Acids and Bases: (is depend to the amount of H^+ or OH^- produced per mole of substance dissolved)

Note: Strong Acids And Bases Dissociate Completely (100%) In Water Most be know them

Stong Acids		Strong Bases	
Hydrobromic acid	HBr	Barium hydroxide	$Ba(OH)_2$
Hydrochloric acid	HCl	Calsium hydroxide	$Ca(OH)_2$
Hydroiodic acid	HI	Lithium hydroxide	$LiOH$
Nitric acid	HNO_3	Potassium hydroxide	KOH
Perchloric acid	$HClO_4$	Sodium hydroxide	$NaOH$
Sulfuric acid	H_2SO_4	Strontium hydroxide	$Sr(OH)_2$

Second Lecture

Equilibrium

Neutralisation

K_w

pH

Calculation of pH

Ionic product constant of water K_w

Pure water is very slightly ionised and has little ability for electrical conductivity according to conductivity experiments done by **Kolrawsh** in 1894. Thus, water is slightly ionised into:

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

The equilibrium constant is K_{eq}

$$K_{eq}= \frac{[\text{products}]}{[\text{reactants}]} = \frac{\alpha_{H^+} \,.\, \alpha_{OH^-}}{\alpha_{H_2O}} \;\ldots\ldots\ldots\ldots(1)$$

$\alpha_x = f_x \,.\, [x]$ ....(2) by Substitute in (1)

$$K_{eq}= \frac{f_{H^+}.[H^+] \times f_{OH^-}.[OH^-]}{f_{H_2O}.[H_2O]} \;\ldots\ldots\ldots\ldots(3)$$

when the solution is very dilute f=1

Ionic product constant of water K_w

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

$K_{eq} = \dfrac{[H^+] \times [OH^-]}{[H_2O]}$ ....(4) very small of water is ionised, H_2O is constant

After that the K_{eq} transform to K_w

$$K_w = [H^+] \times [OH^-] \text{ ....(5)}$$

Where Kw is ionic product constant of water which equals at 25 ℃

$$\mathbf{10.01 \times 10^{-14}}$$

NOTE: The K_w was increases with in increase the temperatures

Second Lecture

Equilibrium

Neutralisation

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Calculation of pH

Ionic product constant of water K_w

$$K_w = [H^+] \times [OH^-] \ldots\ldots\ldots(5)$$

$$10.01 \times 10^{-14} = [H^+] \times [OH^-]$$

At equilibrium of water ionization : $[H^+] = [OH^-]$

$$[H^+] = [OH^-] = \sqrt{10.01 \times 10{-}14} = \mathbf{10^{-7}}$$

When $[H^+] > 10^{-7}$ **the solution is acidic,**

When $[H^+] = 10^{-7}$ **the solution is neutral,**

When $[H^+] < 10^{-7}$ **the solution is basic.**

Also from eq (5)

$$[H^+] = \frac{K_w}{[OH^-]} \quad \text{and} \quad [OH^-] = \frac{K_w}{[H^+]}$$

Hydrogen-Ion Exponent (pH):

It is preferred to express hydrogen ion concentration as ion exponent or *p-function* of hydrogen ion (pH), instead of using negative exponents:

$$pH = -\log[H^+] = \log \frac{1}{[H^+]} \text{ or } [H^+] = 10^{-PH}$$

$$pOH = -\log[OH^-] = \log \frac{1}{[OH^-]} \text{ or } [OH^-] = 10^{-POH}$$

Therefore from equation (5): $[H^+]\,[OH^-] = K_w \times \log$

$$\mp \log[H^+] \mp \log[OH^-] = \mp \log K_w$$

$pH + pOH = pK_w = 14$

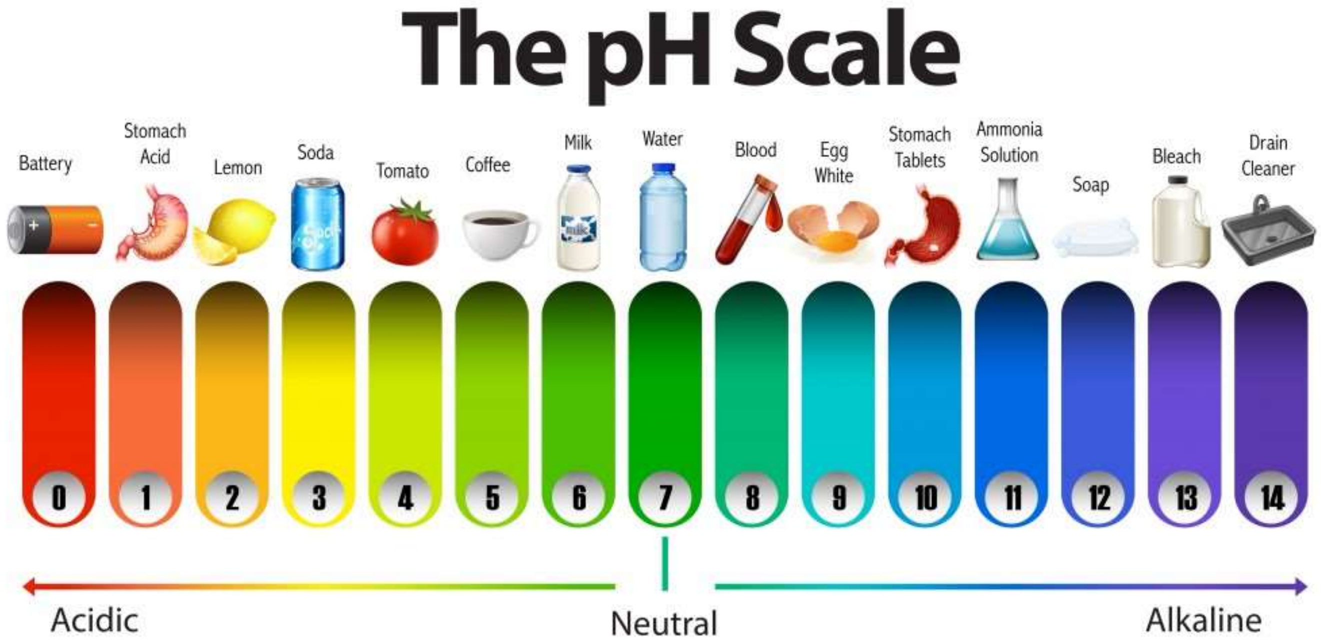
$pH = 14 - pOH$

$pOH = 14 - pH$

$pH = pOH = 7$

Hydrogen-Ion Exponent (pH):

When **pH = 7, the solution is neutral,**
When **pH < 7, the solution is acidic,**
When **pH > 7, the solution is basic.**
Therefore, the pH range is from zero to 14



Second Lecture

Equilibrium

Neutralisation

K_w

pH

Calculation of pH

Calculation of pH

- **Strong Acids**

They are completely ionized such as HCL,HNO_3 and H_2SO_4

Ex. Calculate the pH of 0.01 M of HCl and calculate its pOH ?

Solution:

$$\underset{0.01M}{HCl} \rightarrow \underset{0.01M}{H^+} + \underset{0.01M}{Cl^-}$$

$[H^+] = 0.01M = 10^{-2}$

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

$pOH = 14 – pH$

$pOH = 14 – 2 = 12$ **pOH = 12**

Second Lecture

Equilibrium

Neutralisation

K_w

pH

Calculation of pH

Calculation of pH

- **Strong Acids**

They are completely ionized such as HCL,HNO_3 and H_2SO_4

Ex. What is the pH and pOH of a 2.5 x 10^{-2} M solution of H_2SO_4

Solution:

$$H_2SO4 \rightarrow 2H^+ \quad + \quad SO_4^{-2}$$

2.5 x 10^{-2} 2.5 x 10^{-2} 2.5 x 10^{-2}

$[H^+] = 2\times 2.5 \times 10^{-2}$

$pH = -\log[2\times 2.5 \times 10^{-2}] = -\log 5.0 \times 10^{-2} = -0.7+2 = 1.3$

pH =1.3

pOH = 14 – pH

pOH = 14 – 1.3 = 12.7

pOH = 12.7

Second Lecture

Equilibrium

Neutralisation

K_w

pH

Calculation of pH

Calculation of pH

- **Strong bases**

They are completely ionized such as NaOH, KOH and $Ba(OH)_2$.

Ex. Calculate the pOH of 0.01 M of NaOH and calculate its pH ?

Solution:

$$\underset{0.01M}{NaOH} \rightarrow \underset{0.01M}{Na^+} + \underset{0.01M}{OH^-}$$

$[OH^-] = 0.01M = 10^{-2}$

$pOH = -\log[OH^-] = -\log 10^{-2} = 2$ **pOH = 2**

$pH = 14 – pOH$

$pH = 14 – 2 = 12$ **pH = 12**

Second Lecture

Equilibrium

Neutralisation

K_w

pH

Calculation of pH

Calculation of pH

- **Strong bases**

They are completely ionized such as NaOH, KOH and $Ba(OH)_2$.

Ex. Calculate the pOH of 0.03 M of $Ba(OH)_2$ and calculate its pH ?

Solution:

$$\underset{0.03M}{Ba(OH)_2} \rightarrow \underset{0.03M}{Ba^{+}} + \underset{0.03M}{2OH^{-}}$$

$[OH^-] = 2 \times 0.03M = 6\times10^{-2}$

$pOH = -\log[6\times10^{-2}] = -0.77 + 2$ **pOH = 1.23**

$pH = 14 - pOH$

$pH = 14 - 1.23 = 12$ **pH = 12.77**

Second Lecture

REVIEW

- What's is the un equilibrium in chemistry ?
- Neutralization reaction
- The strength of an acid and a base in a solution depends on what ?
- Derive an equation of Ionic product constant of water K_w
- Express of Hydrogen-Ion Exponent (pH)
- pH scale and know the difference between pH & pOH?
- Exercises for calculation the pH of strong acid & base

Home Work

1. Calculate the pH and pOH of the following acidic solutions?
   (a) 0.082 M HNO_3
   (b) 0.5 mL of 0.1M of HCl add to 50 ml water

2. Calculate the pH and pOH of the following basic solutions?
   (a) 0.055 M KOH
   (b) 0.8 mL of 0.2M of NaOH add to 50 ml water