



ENGINEERING CHEMISTRY-II

As per the latest syllabus of Diploma in Engineering Courses Under
Jharkhand University of Technology, Ranchi

Dr. Altaf Ahmad

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Published by :

Romesh Chander Khanna & Vineet Khanna
for **KHANNA PUBLISHERS**
2-B, Nath Market, Nai Sarak
Delhi- 110 006 (India)

Website : www.khannapublishers.in

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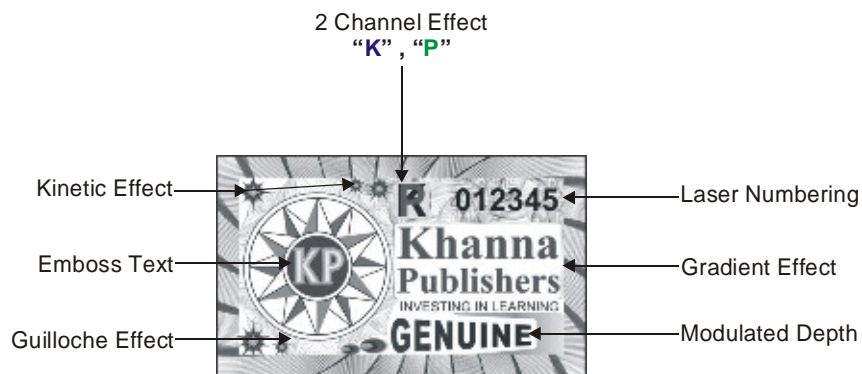
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ISBN No. 978-93-87394-74-2

First Edition : 2020

Preface to the First Edition

This book is exclusively written for the students of various branches of diploma in engineering throughout the country in general and for the students of Jharkhand in particular, keeping in view of their professional requirements, after entering into their practical life.

The objective of the present book is to serve the students with an elementary knowledge of chemistry. The syllabi of chemistry taught in the name of engineering chemistry in different diploma colleges is fairly diverse in nature. It is therefore quite impossible to provide a complete coverage at all the topics in a limited space for all students. However, author have dealt with modern views of the topics of the syllabi and attempted to give a major coverage at the recent syllabi taught in various diploma level institutions.

This book embodies eight chapters which are strictly as per the syllabus prescribed by the Jharkhand University of Technology, Jharkhand for the students of second semester of three years Diploma in Engineering. Each chapter consists of a methodical introduction, historical background, discussion of basic physico-chemical principles involved and practical applications and significance. Chapters on Water Treatment also contain systematic methods of solving problems followed by several worked out examples. Further, at the end of each chapter various review questions are given which also include multiple choice questions and answers.

This book is written solely with a conviction to cater the academic and professional requirements of the students of all branches of Diploma in Engineering.

Any suggestions and constructive criticism towards this objective are welcome.

December 2019

Dr. Altaf Ahmad

Aknowledgement

I take this opportunity to express my profound gratitude and deep regards to Dr. M. K. Banerjee (Retired Senior Scientist, National Metallurgical Laboratory, Jamshedpur and Retired Professor and Head, Department of Chemistry, NIT, Jamshedpur.) for his constant encouragement, support and critical discussion throughout the writing of the manuscript of this book.

I would like to thanks the Principal, Al-Kabir Polytechnic, Jamshedpur, for his valuable advice and moral support.

I whole heartedly express my thanks to all my colleagues at AL-Kabir Polytechnic, Jamshedpur particularly to Mrs. Sanobar and Mr. Naseem for their co-operation in proof reading, valuable advice and moral support.

I also express my sincere thanks to all concerned faculty and staff of Al-Kabir Polytechnic, Jamshedpur who contributed in my experimental work and completing this thesis.

I also extent my sincere thanks to all concerned faculty and staff of National Institute of Foundry and Forge Technology (NIFFT), Hatia, Ranchi, who contributed in any way in completing this research work.

I also thank to my wife Shama for her love, moral support & patience that has given me, a boost and help to overcome many difficult times. I also thank my daughter Aalia and son Mujtab, whom I believe to be the most sufferer during the period of writing this book, because I couldn't spare as much time as they needed.

Last but not the least, I extend my thanks to family members and especially to my father for their unconditional support, sacrifices, understanding, patience, constant pray and motivational spirit, which helped me execute my work smoothly.

Dr. Altaf Ahmad

Dedicated to my Father-in-law

With Veneration and Reverence

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Electrochemistry

Syllabus: Conductivity of Electrolytes – Concept of Ohm’s Law, Specific Conductance, Specific Resistance, Equivalent Conductivity and Molar Conductance, Variation of Specific, Molar and Equivalent Conductance with dilution. Concept of: Cell Constant, pH, pOH and Buffer solution. Numerical based on pH and pOH. Application of pH and Buffer solution.

1.1. INTRODUCTION

Electrochemistry is the branch of chemistry which deals with the interactions of electrical energy with chemical species. Broadly it is divided into two categories. These are: (i) occurrence of chemical reactions at the expense of electrical energy, and (ii) transformation of chemical energy into electrical energy *i.e.*, generation of electrical energy by spontaneous redox reactions. The electrolysis (electroplating, electro-refining, electro-typing etc.) involves the occurrence of chemical reactions at the expense of electrical energy) whereas, chemical energy is transformed into electrical energy in galvanic cells, fuel cells and batteries. All electrochemical reactions involve transfer of electrons and are therefore, called oxidation-reduction (redox) reactions.

1.2. CONDUCTORS

Substances that give passage to the flow of electric current through them are called electrical conductors or simply conductors. For example, metals, graphite, fused salts, aqueous solutions of acids, bases and salts. Conductors are of two types.

1.2.1. Electronic conductors

Substances that give passage to the flow of electric current because of the presence of free electrons are called electronic conductors *i.e.*, these conductors transfer electric current by transfer of electrons, without the transference of any matter. Metals such as aluminium, copper, silver etc., graphite (an allotropic form of carbon) and various alloys belong to this class. These materials contain electrons which are relatively free to move. Silver is the best known electronic conductor. Graphite is the only non-metal which is electronic conductor. All metals conduct electric current because of the presence of free electrons hence, electronic conductors are sometimes also referred to as metallic conductors.

1.2.2. Electrolytic conductors

Substance that give passage to the flow of electric current because of the presence of free ions are called electrolytic conductors *i.e.*, these conductors transfer electric current by transfer of ions (transference of matter). Conductors like aqueous solutions of acids, bases and salts in which the flow of electric current is accompanied by chemical decomposition are known as electrolytic conductors.

1.3. INSULATORS

Substances which do not give passage to the flow of electric current through them are called electrical insulators or simply insulators. For example: wood, glass, rubber, plastics etc.

Distinction between Electronic and Electrolytic Conductor

<i>Electronic conductors</i>	<i>Electrolytic conductors</i>
1. These conductors conduct electric current by movement of electrons.	1. These conductors conduct electric current by movement of ions.
2. In this transference of matter is not involved.	3. In this transference of matter is involved.
4. In this no chemical reaction takes place.	2. In this chemical reaction takes place, <i>i.e.</i> , ions are oxidised or reduced at the electrodes.
3. Resistance of the conductor increases with increase in temperature.	4. Resistance of the conductor decreases with increase in temperature.
5. In this Faraday's law is not followed.	6. In this Faraday's law is followed.

1.4. ELECTROLYTE

Substances that dissociate into their constituent ions either in their aqueous solutions or in their molten states and hence conduct electricity are called electrolytes. For example: HCl, H₂SO₄, CH₃COOH, NaOH, NH₄OH, NaCl, KNO₃, NH₄Cl etc.

1.5. CLASSIFICATION OF ELECTROLYTES

The capacity of an electrolyte to conduct electric current is measured in terms of degree of ionization (α), which is different for different electrolytes. The degree of ionization may be defined as the number of moles of electrolyte dissociated into ions per mole of the electrolyte. Based upon the extent of ionization, electrolytes are classified into following two classes.

1.5.1. Strong Electrolytes

Substances that dissociate almost completely into their constituent ions either in their aqueous solutions or in their molten states are called strong electrolytes. For example: strong acids such as, HCl, HNO₃, H₂SO₄ etc., strong bases such as NaOH, KOH etc., salts of strong acids and strong bases such as NaCl, KNO₃, K₂SO₄ etc.

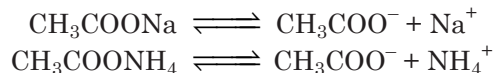
For strong electrolytes, $\alpha = 1$

1.5.2. Weak Electrolytes

Substances that do not dissociate completely into their constituent ions either in their aqueous solutions or in their molten states are called weak electrolytes. For example: weak acids such as, CH₃COOH (acetic acid), C₆H₅COOH (benzoic acid), H₃BO₃ (boric acid), weak bases such as, CH₃NH₂ (methanamine), NH₄OH (ammonium hydroxide) salts of weak acids and weak bases CH₃COONa (sodium acetate), CH₃COONH₄ (ammonium acetate) etc.

For weak electrolytes, $\alpha < 1$

As weak electrolytes do not ionised completely in their solutions, there exist an equilibrium between ions and undissociated molecules of electrolyte.



1.6. NON-ELECTROLYTE

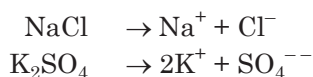
Substances that do not dissociate into their constituent ions either in their aqueous solutions or in their molten states and do not conduct electricity are called non-electrolytes. For example: solutions of cane sugar, glycerine, acetone etc.

For non-electrolytes, $\alpha = 0$

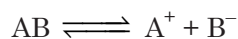
1.7. ARRHENIUS THEORY OF ELECTROLYTIC DISSOCIATION

In order to explain the properties of electrolytic solutions, Arrhenius forwarded a comprehensive theory in 1884, called as theory of electrolytic dissociation or ionic theory. Following are the important postulates of the theory.

1. The theory assumes that the solid electrolytes are composed of ions which are held together by strong electrostatic forces of attraction.
2. When an electrolyte is dissolved in water (solvent) it dissociates into their constituent ions *i.e.*, cations and anions.



3. These constituent ions of the electrolyte remain hydrated (solvated) in the solution.
4. The process of splitting of electrolyte molecule into constituent ions is called ionization.
5. The electrolytic solution always remains electrically neutral as the total charges on cations are always equal to the total charges on anions.
6. When an electric current is passed through the solution, the cations move towards cathode and anion moves towards anode.
7. It is observed that all electrolytes do not ionize to the same extent. Some are almost completely ionised while some are only feebly ionised.
8. Ions present in solution constantly reunite to form neutral electrolyte molecules and thus there exist a state of dynamic equilibrium between ions and the neutral molecules.



9. The properties of electrolytes in the solution are the properties of ions present in it.
10. The ions act like molecule towards colligative properties.
11. The conductivity of the electrolytic solution depends upon the number and nature of ions present in the solution.

1.8. OHM'S LAW

The law was given by a German physicist 'George Simon Ohm' in 1827. The law states that at constant physical conditions (temperature, pressure, length of the conductor etc.) the current following through a conductor is directly proportional to the potential difference across the ends of the conductor.

i.e.,

$$I \propto V$$

or

$$I = \frac{V}{R}$$

or

$$V = IR \quad \dots(1.1)$$

where, R is a constant called resistance.

Equation (1.1) is a straight line equation of the form $y = mx + c$ (where, ' c ' is the intercept and equals to zero and ' m ' is the slope). Thus, a plot of V vs. I will be straight line passing through origin (intercept equals zero) with slope equals R , as shown below.

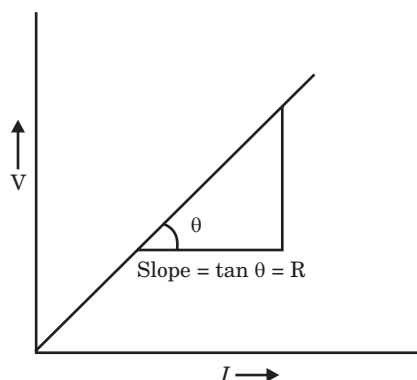


Fig. 1.1: Graphical representation of Ohm's law

1.9. RESISTANCE (R)

The electrical resistance is the property by which it opposes or obstructs the flow of current in an electrical circuit. In other words, resistance is the hindrance to the flow of electrons. It is measured in ohms and symbolized by the Greek letter omega (Ω).

The resistance of a conductor is directly proportional to the length of the conductor and inversely proportional to the area of cross-section of the conductor.

Let ' l ' be the length and ' a ' be the area of cross-section of a conductor.

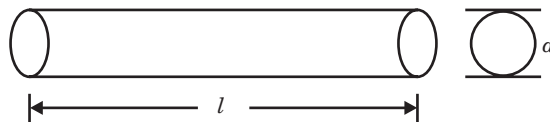


Fig. 1.2: Dependence of Resistance

Then,

$$R \propto l$$

and

$$R \propto \frac{1}{a}$$

or

$$R \propto \frac{l}{a}$$

or

$$R = \rho \frac{l}{a} \quad \dots(1.2)$$

where, ρ (rho) is a constant called specific resistance or resistivity of the conductor.

In an electrochemical cell ' l ' is the distance between two electrodes and ' a ' is the area of cross-section of each electrode.

1.10. FACTORS AFFECTING RESISTANCE OF A CONDUCTOR

The resistance of a conductor is generally influenced by the following factors:

1. Nature of material : Conductors have low electrical resistance and low resistivity. On the other hand insulators have high electrical resistance and high resistivity. Whereas the resistance of semi-conductors lie between conductors and insulators.

2. Length of the conductor : The resistance of a conductor is directly proportional to the length of the conductor. Thus, the resistance increases with increase in the length of the conductor.

3. Area of cross-section of the conductor : The resistance of a conductor is inversely

proportional to the area of cross-section of the conductor. Thus, the resistance of thin wire is more than a thick wire.

4. Temperature : In general the resistance of a conductor increases with temperature *i.e.*, resistance of a conductor is directly proportional to the temperature.

1.11. CELL CONSTANT (G^*)

The ratio of the distance between electrodes to the area of cross-section of electrodes is called cell constant. In other words the ratio ' l/a ' is called cell constant. It is measured in per cm (cm^{-1}) and is denoted by G^* .

Thus, $G^* = l/a$... (1.3)

where, l = distance between two electrodes in cm

and a = area of cross-section of each electrode in cm^2

Since, it is not possible to measure both ' l ' and ' a ' with maximum accuracy, the value of cell constant cannot be determined directly with precision. Therefore the cell constant is determined indirectly by measuring the conductance of a standard (say N/10) KCl solution in a conductivity cell whose specific conductance is known. From these data the cell constant can be calculated.

The value of cell constant of a particular cell is fixed and does not change unless the electrodes are changed.

Also, Resistance = Specific resistance \times Cell constant.

1.12. CONDUCTANCE (G or C)

The ease with which a conductor allows the current to pass through it is called conductance. It is the ability of a system to conduct electricity. It is measured in ohm^{-1} or mho or Siemens (S).

The conductance of a conductor is given as the reciprocal of the resistance.

i.e., $C = 1/R$... (1.4)

or $C = \frac{1}{\rho} \times \frac{a}{l}$

or $C = \kappa \times \frac{a}{l}$... (1.5)

where, ' κ ' (kappa) is a constant called specific conductance or conductivity.

Also, Conductance = $\frac{\text{Specific conductance}}{\text{Cell constant}}$... (1.6)

1.13. SPECIFIC RESISTANCE or RESISTIVITY (ρ)

We have, $R = \rho \frac{l}{a}$

If, $l = 1 \text{ cm}$

and $a = 1 \text{ cm}^2$

Then $R = \rho$

Thus, the specific resistance of an electrolytic solution is the resistance offered when the electrodes having area of cross-section 1 cm^2 are apart from each other by 1 cm. It is measured in ohm cm. The specific resistance is given as ;

$$\rho = R \frac{a}{l} \quad \dots (1.7)$$

$$\text{Also, Specific resistance} = \frac{\text{Resistance}}{\text{Cell constant}} \quad \dots(1.8)$$

1.14. SPECIFIC CONDUCTANCE or CONDUCTIVITY (κ or σ)

$$\text{We have,} \quad C = \kappa \times \frac{a}{l}$$

$$\text{If,} \quad l = 1 \text{ cm}$$

$$\text{and} \quad a = 1 \text{ cm}^2$$

$$\text{Then} \quad C = \kappa$$

Thus, the specific conductance of an electrolytic solution is the conductance offered when the electrodes having area of cross-section 1 cm^2 are apart from each other by 1 cm . The specific conductance of an electrolytic solution is a measure of its ability to conduct electricity. It is measured in $\text{ohm}^{-1} \text{ cm}^{-1}$ or mho cm^{-1} or s cm^{-1} and is denoted by ' κ ' (kappa).

The specific conductance is given as the reciprocal of specific resistance.

$$\text{i.e.,} \quad \kappa = \frac{1}{\rho}$$

$$\kappa = \frac{1}{R(a/l)}$$

$$\kappa = C \frac{l}{a} \quad \dots(1.9)$$

$$\text{Also, Specific conductance} = \text{Conductance} \times \text{Cell constant}$$

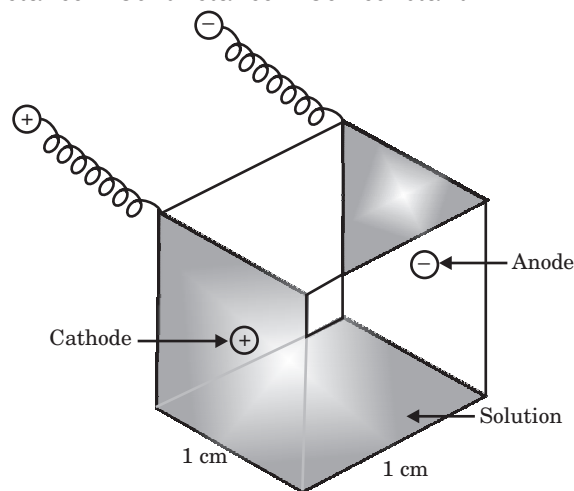


Fig. 1.3: Representation of specific conductance

1.15. MOLAR CONDUCTANCE (Λ_m)

The conductance offered by all the ions produced by the ionization of 1 mole of an electrolyte present in a solution is called molar conductance or molar conductivity. The molar conductance is given as;

$$\Lambda_m = \kappa \times V \quad \dots(1.10)$$

where,

$$\Lambda_m = \text{Molar conductance}$$

$$\kappa = \text{Specific conductance}$$

V = Volume of solution in ml

If ' M ' be the molarity (number of gram moles of solute dissolve in one litre solution) of the solution then,

$$V = \frac{1000}{M}$$

Therefore, $\Lambda_m = \kappa \times \frac{1000}{M}$... (1.11)

The unit of molar conductance is $\text{ohm}^{-1} \text{cm}^2 \text{mole}^{-1}$ or $\text{mho cm}^2 \text{mole}^{-1}$ or $\text{s cm}^2 \text{mole}^{-1}$.

1.16. EQUIVALENT CONDUCTANCE (Λ_{eq})

The conductance offered by all the ions produced by the ionization of 1 gram equivalent of an electrolyte present in a solution is called equivalent conductance or equivalent conductivity. The equivalent conductance is given as;

$$\Lambda_{eq} = \kappa \times V \quad \dots (1.12)$$

where,

Λ_{eq} = Equivalent conductance

κ = Specific conductance

V = Volume of solution in ml.

If ' N ' be the normality (number of gram equivalent of solute dissolve in one litre solution) of the solution then,

$$V = \frac{1000}{N}$$

Therefore, $\Lambda_{eq} = \kappa \times \frac{1000}{N}$... (1.13)

The unit of molar conductance is $\text{ohm}^{-1} \text{cm}^2 \text{eq}^{-1}$ or $\text{mho cm}^2 \text{eq}^{-1}$ or $\text{s cm}^2 \text{eq}^{-1}$.

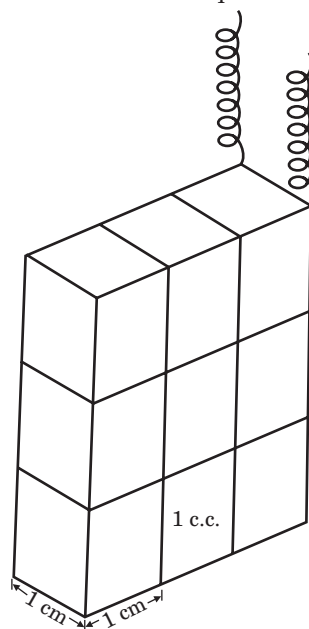


Fig. 1.4: Representation of molar and equivalent conductance

1.17. VARIATION OF CONDUCTIVITIES WITH DILUTION

The conductance of a solution mainly depends upon two factors.

- (i) Number of ions present in the solution, and
- (ii) Ionic mobility, which is the speed of the ions under unit potential gradient.

Either the increase in number of ions in solution or the increase in ionic mobility, will increase the conductance of the solution. Following are the effect of dilution on different conductivities.

1.18. EFFECT OF DILUTION ON SPECIFIC CONDUCTANCE

Specific conductance is the conductivity of unit volume of solution. Thus the specific conductance (κ) of a solution increases with increase in concentration of solution as the number of ions per unit volume increases. In other words the specific conductance (κ) decreases with dilution (decrease in concentration), because number of ions in unit volume of solution decreases with dilution.

1.19. EFFECT OF DILUTION ON MOLAR CONDUCTANCE AND EQUIVALENT CONDUCTANCE

Both the Equivalent conductance (Λ_{eq}) and Molar conductance (Λ_m) increase with dilution, because with dilution both the number of ions and ionic mobility increases. Strong and weak electrolytes have different degree of ionization and therefore they behave differently in their solutions. The conductance behaviour and effect of concentration on conductance is different for both strong and weak electrolytes. Let us understand the conductance behaviour of two types of electrolytes separately.

1.19.1. Conductance Behaviour of Strong Electrolytes

Both the Molar conductance and Equivalent conductance of strong electrolytes shows only a small decrease with increase in concentration. In other words, the molar and equivalent conductance of strong electrolytes increases only slightly with dilution (decrease in concentration).

In fact, the degree of ionization of strong electrolyte is not affected by dilution and they are completely dissociated into their constituent ions in solution at ordinary dilutions. However, at higher concentrations, these dissociated ions are close to each other and thus, the inter-ionic attractions are greater. This inter-ionic attraction retard the motion of the ions and thus, conductivity is low. With dilution (decrease in concentration), the inter-ionic attraction decreases because the ions move away from each other. This results in the increase in molar and equivalent conductance with dilution and becomes maximum at infinite dilution.

The Molar conductance of strong electrolyte is given by Debye-Huckel-Onsagar equation as follows.

$$\Lambda_m = \Lambda_m^\infty - b\sqrt{c} \quad \dots(1.14)$$

where, Λ_m = Molar conductance at 'V' dilution

Λ_m^∞ = Limiting Molar conductance *i.e.*, Molar conductance at infinite dilution.

c = Concentration of solution

b = It is a constant that depend only on known quantities such as temperature, the charges on the ions and the dielectric constant and viscosity of the solvent.

1.19.2. Conductance Behaviour of Weak Electrolytes

Weak electrolytes dissociate into ions to a lesser extent and thus, furnish less number of

ions for conductance as compared to strong electrolytes. As a result the conductance of a weak electrolyte is much less than that of a strong electrolyte at same concentration.

In general the molar and equivalent conductance of weak electrolytes increases largely with dilution (decrease in concentration) especially near infinite dilution. This is due to the fact that with dilution the ionization of weak electrolyte and also the ionic mobility increases. This results in more number of ions in solution and thus, there is an increase in Molar and Equivalent conductance. However, the conductance of a weak electrolyte never approaches to a limiting value.

Fig. 1.5 shows the effect of dilution on Molar conductance of some strong and weak electrolytes.

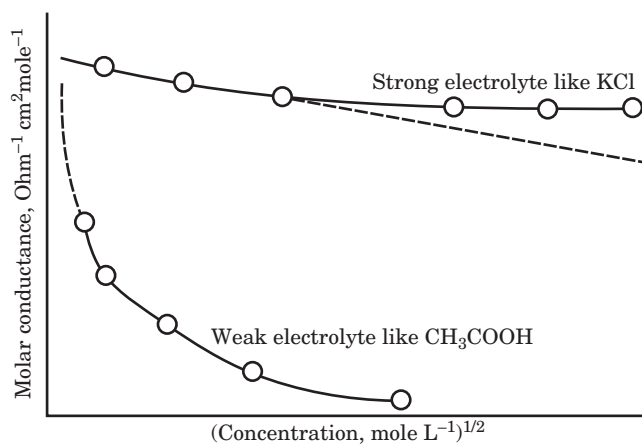


Fig. 1.5: Variation of Molar conductance with dilution

Fig. 1.5 reveals that strong electrolytes like KCl have high value of molar conductance even at low concentration. Their molar conductance increases slowly with dilution *i.e.* there is no rapid increase in their molar conductance with dilution. In the case of strong electrolytes, there is a tendency for molar conductance to approach a limiting value when the concentration approaches zero. When the whole of the electrolyte has ionised, further addition of the water does not bring any change in the value of equivalent conductance. This stage is called infinite dilution. The equivalent conductance has a limiting value at infinite dilution and is represented by Λ_m^∞ . A plot between the molar conductance and \sqrt{c} (where c is the concentration) is a straight line having y -intercept equal to Λ_m^∞ (molar conductance at infinite dilution).

Weak electrolytes like acetic acid (CH_3COOH) have a low value of molar conductance at higher concentration and there is a rapid increase in the value of molar conductance with dilution. There is no indication that a limiting value of molar conductance can be attained even when the concentration approaches zero (infinite dilution). A plot between the molar conductance and \sqrt{c} (where c is the concentration) is not a straight line. Weak electrolytes have lower molar conductivities and lower degree of dissociation at higher concentrations which increases steeply at lower concentrations. Therefore, limiting molar conductance *i.e.*, molar conductance at infinite dilution, Λ_m^∞ , cannot be obtained by extrapolation of molar conductance at zero concentration. The molar conductance at infinite dilution for weak electrolytes can be obtained using Kohlrausch law of independent migration of ion.

Thus it can be concluded that the molar conductance of electrolytes whether strong or weak, increases with dilution and reaches a maximum or limiting value at infinite dilution.

In the case of strong electrolytes this limiting value can be obtained by extrapolation of the graph of molar conductance to zero concentration but in the case of weak electrolytes it cannot be obtained accurately. An indirect method for obtaining Λ_m^∞ for weak electrolytes has been given by Kohlrausch. The values of molar conductances of some electrolytes at 298K are given in Table 1.1.

Table 1.1: Molar conductances, ($\text{Ohm}^{-1} \text{cm}^2 \text{mole}^{-1}$) of some electrolytes at 298K

Concentration	Strong electrolytes			Weak electrolytes	
	HCl	NaOH	NaCl	CH ₃ COOH	NH ₄ OH
0.1 M	391.32	183.0	106.74	5.2	3
0.05 M	399.09	192.0	111.06	7.4	11.3
0.02 M	407.24	201.0	115.76	11.6	34.0
0.01 M	412.00	238.0	118.51	22.8	46.9
0.005 M	415.80	240.8	120.65	48.6	—
0.001 M	421.40	244.7	123.74	—	—
0.0005 M	422.74	245.6	124.50	—	—

1.20. IONIC PRODUCT OF WATER

Water is a weak electrolyte and ionizes according to the equation



Applying law of mass action at the state of equilibrium, the value of dissociation constant, K_c can be given as

$$K_c = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

Since water is a weak electrolyte and ionizes very feebly, hence the concentration of unionised water molecules, $[\text{H}_2\text{O}]$, remains practically constant. Thus the product of K_c and $[\text{H}_2\text{O}]$ will also be a constant and designated as K_w . So,

$$K_w = [\text{H}^+][\text{OH}^-] \quad \dots(1.16)$$

where, K_w is called ionic product of water and is defined as the product of the concentrations of H^+ and OH^- ions present in water at a particular temperature. The value of ionic product of water increases with increase in temperature, i.e., the concentration of H^+ and OH^- ions increases with the increase of temperature.

Experimentally, the ionic product of water was found to be $1.0 \times 10^{-14} \text{mole}^2 \text{litre}^{-2}$ at 298K (25°C). Thus,

$$[\text{H}^+][\text{OH}^-] = 10^{-14} \quad \dots(1.17)$$

From equation (1.11) we have

$$[\text{H}^+] = [\text{OH}^-]$$

therefore,

$$[\text{H}^+][\text{H}^+] = 10^{-14}$$

or

$$[\text{H}^+] = 10^{-7}$$

Taking negative logarithm both sides we get,

$$-\log [\text{H}^+] = -\log 10^{-7}$$

or

$$\text{pH} = 7$$

Therefore pH of pure water is 7.

Similarly we can have,

$$[\text{OH}^-] [\text{OH}^-] = 10^{-14}$$

or
$$[\text{OH}^-] = 10^{-7}$$

Taking negative logarithm both sides we get,

$$-\log [\text{OH}^-] = -\log 10^{-7}$$

or
$$\text{pOH} = 7$$

Therefore pOH of pure water is also 7.

Thus, for pure water which is neutral in nature,

$$\text{pH} = \text{pOH} = 7$$

and
$$[\text{H}^+] = [\text{OH}^-] = 10^{-7} \text{ mole litre}^{-1}.$$

This shows that at 25°C (298K), in 1 litre (approximately 55.5 moles) of water, only 10^{-7} moles of water remain in ionic form. When an acid or a base is added to water, the ionic product, $[\text{H}^+][\text{OH}^-]$, remains constant, *i.e.*, equal to K_w but concentrations of H^+ and OH^- ions do not remain equal. The addition of acid increases the hydrogen ion concentration and addition of a base increases the hydroxyl ion concentration.

Therefore for a solution,

$$[\text{H}^+] [\text{OH}^-] = 10^{-14} \text{ mole}^2 \text{ litre}^{-2} \quad \dots(1.18)$$

and
$$\text{pH} + \text{pOH} = 14 \quad \dots(1.19)$$

1.21. pH

The negative logarithm of hydrogen ion concentration present in a solution expressed in mole litre⁻¹ is called pH. The symbol has been taken from Danish word “Potenz de hydrogen ion.”

$$\text{pH} = -\log [\text{H}^+] \quad \dots(1.20)$$

or
$$\text{pH} = \log \frac{1}{[\text{H}^+]} \quad \dots(1.21)$$

If the concentration of H^+ ion in a solution be 10^{-x} mole litre⁻¹, then

$$\text{pH} = -\log [\text{H}^+]$$

or
$$\text{pH} = -\log 10^{-x}$$

or
$$\text{pH} = -(-x) \log 10$$

or
$$\text{pH} = x$$

If, $x < 7$ solution will be acidic

If, $x > 7$ solution will be alkaline

If, $x = 7$ solution will be neutral

1.22. pOH

The negative logarithm of hydroxyl ion concentration present in a solution expressed in mole litre⁻¹ is called pOH.

$$\text{pOH} = -\log [\text{OH}^-] \quad \dots(1.22)$$

or,
$$\text{pOH} = \log \frac{1}{[\text{OH}^-]} \quad \dots(1.23)$$

If the concentration of OH^- ion in a solution be 10^{-x} mole litre⁻¹, then

$$\begin{aligned} & \text{pOH} = -\log [\text{OH}^-] \\ \text{or} & \text{pOH} = -\log 10^{-x} \\ \text{or} & \text{pOH} = -(-x) \log 10 \\ \text{or} & \text{pOH} = x \end{aligned}$$

If, $x < 7$ solution will be alkaline

If, $x > 7$ solution will be acidic

If, $x = 7$ solution will be neutral

1.23. pH SCALE

The pH scale also called as Sorensen scale was invented by a Danish Chemist Soren Sorensen to measure the acidity of beer in brewery. The pH scale measure the concentration of hydrogen ions in solution. The more the hydrogen ions, the stronger will be the acid. Now-a-days the pH scale is used to express the acidity or alkalinity of a solution, The pH scale can be shown as below.

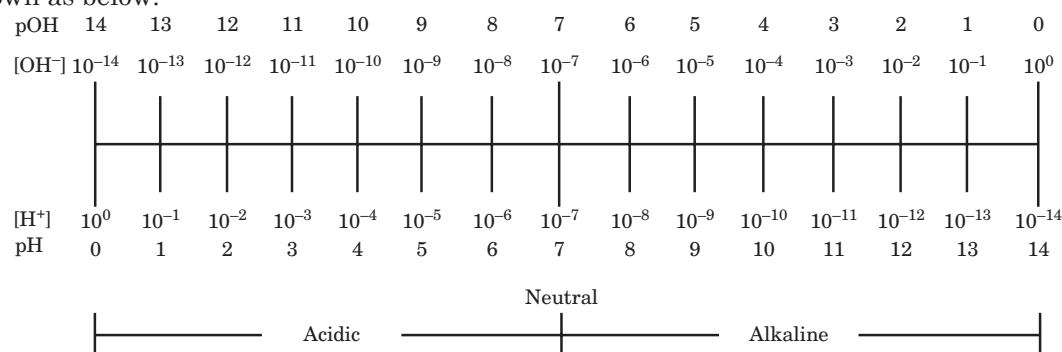


Fig. 1.6. pH scale

A pH scale consists of numbers from 1 to 14, with 7 as the middle (neutral) point. Values below 7 indicate acidity which increases as the number decreases, 1 being the most acidic. Values above 7 indicate alkalinity which increases as the number increases, 14 being the most alkaline. The pH scale, however, is not a linear scale like a centimetre or inch scale (in which two adjacent values have the same difference). It is a logarithmic scale in which two adjacent values increase or decrease by a factor of 10. For example, a pH of 2 is 10 times more acidic than a pH of 3 and 100 times more acidic than a pH of 4. In a similar way, a pH of 10 is 10 times more alkaline than a pH of 9 and 100 times more alkaline than a pH of 8.

If the pH of a solution is decreased by one unit, the concentration of H⁺ ions in the solution is increased 10 times. For example, when pH of a solution is decreased from 6 to 3, the H⁺ ion concentration is increased $10 \times 10 \times 10 = 1000$ times, because for each unit of pH decreased, the H⁺ ion concentration is increased by 10 times. Similarly, if the pH of a solution is increased from 4 to 5, the concentration of H⁺ ion in the solution is decreased by 10 times. For example, when pH of a solution is increased from 5 to 7, the H⁺ ion concentration is decreased $10 \times 10 = 100$ times, because for each unit of pH increased, the H⁺ ion concentration is decreased by 10 times. Thus a lower value of pH indicates that the solution is more acidic and a higher value of pH indicates that the solution is more basic. To conclude, with increase in the value of pH, the concentration of H⁺ ion in the solution decreases and thus the acidic strength decreases or basic strength increases.

1.23.1. Limitations of pH Scale

1. This is an arbitrary scale.
2. pH values of the solutions do not give an immediate idea of the relative strengths of the solutions. For example, a solution of pH = 1 has a hydrogen ion concentration 100 times that of a solution pH = 3 (not three times). A solution of HCl having concentration 4×10^{-5} M is twice concentrated than that of a solution of HCl having concentration 2×10^{-5} M, but the pH values of these solutions are 4.40 and 4.70 (not double).
3. The pH scale also fails to indicate the pH of very strong acids or bases. For example, the pH value of 1M strong acid such as HCl is zero. If the concentration is 2 M, 3 M, 10 M etc. then the respective pH values will be negative which is not included in pH scale.
4. The pH scale fails to explain the pH of acids or bases having very low concentration *i.e.*, less than 10^{-7} M. A solution of an acid having very low concentration, say 10^{-8} M, cannot have pH 8, as shown by pH formula rather the actual pH value of this solution will be less than 7.
5. The pH values outside the range of pH scale (0 – 14) are possible but do not tend to be accurate because strong acids and bases do not dissociate completely in highly concentrated solutions.
6. pH scale is confined to dilute solutions only.
7. The scale provides no idea about the temperature dependence of pH.

1.24. APPLICATION OF pH

Some of the important applications of pH in our daily life are following:

1. **pH of human body** : Most of the biochemical reactions essential for the survival for life take place in an aqueous environment, however, it is our blood, plasma and interstitial fluids surrounding the cells that are most sensitive to acid-alkaline imbalance. The nature has designed the human body to maintain a very delicate pH balance in its fluids, tissues and systems.
2. **pH in our digestive system**: The hydrochloric acid is produced in our stomach which helps in digestion of food we eat without harming the stomach. But when the amount of acid produced in stomach goes beyond a certain threshold limit, it causes pain and irritation in stomach. In order to neutralize the effect of this excess acid some mild bases are used. These mild bases are called antacids. Magnesium hydroxide (milk of magnesia) and sodium bicarbonate (Baking soda) are two common antacids.
3. **pH change as the cause of tooth decay** : When we eat food containing sugar, then the bacteria present in our mouth break down the sugar to form acids and lactic acid is one such acid. The formation of these acids leads to a decrease in the pH value in the mouth. The white enamel coating on our teeth is of insoluble calcium phosphate, which is quite hard and is not affected by water. But, when the pH in the mouth falls below 5.5, the acid become strong enough to attack the enamel of our teeth and the enamel gets corroded and decay of tooth occurs. As water has a direct access to the roots and decay of teeth occurs. The best way to prevent tooth decay is to clean the mouth thoroughly after eating food. The tooth pastes contain some basic ingredients and they help in neutralising the effect of the acids.
4. **pH in fatigue muscle** : As a result of physical exercise, stiffness and pain in the muscle starts due to the formation of lactic acid. This results in the reduction of pH in the muscle and the supply of oxygen in the muscle is reduced. This causes difficulty in the release of energy leading to increase in the rate of anaerobic metabolism. As a result, lactic acid gets accumulated in the muscles.

5. **pH in self-defense mechanism of animal and plants :** The venom injected by honey bee stings is acidic in nature as it contains formic acid, which causes pain and irritation. Hence, a mild base like baking soda is applied to neutralize the acid and treat the wound. Similarly, nettle leaves, which have stinging hairs, when someone accidentally touches its hairs, a painful effect is produced. When the affected area is rubbed with dock plant which is alkaline in nature neutralizes the acid in the affected area. On the other hand when a wasp stings, it injects an alkaline liquid into the skin that generates malicious effects. This effect can be neutralized by rubbing the affected area (stung area) with vinegar.
6. **pH of soil :** In general soils are acidic in nature. Plants require definite pH range for their proper growth. Many plants do not grow in highly acidic or highly alkaline soil. Therefore, such soils are treated properly to achieve ideal pH for plant growth. For example, highly acidic soil is treated by spreading quick lime or slaked lime or calcium carbonate to lower its acidity.
7. **pH in maintaining brilliance of metals :** Copper vessel gets tarnished because of the formation of basic copper oxide on its surface. When it is rubbed with a piece of lemon, the citric acid present in the lemon reacts with the basic copper oxide and forms copper citrate which is washed away and the shining appearance regains.
8. **pH in swimming pools :** One of the most important steps of swimming pool is water balance. The optimum pH for swimming pool is 7.2 – 7.4, which is also the pH in human eye and mucous membrane. This pH gives good chlorine disinfection. Water with low pH value is regarded as aggressive water, which damages the mechanical components of the pool and causes the irritation of eye and mucous membrane. Whereas, water with high pH value has poor chlorine disinfection and causes skin irritation. Therefore, the pH of water in swimming pool is maintained by using suitable additives.
9. **pH in food preservation:** Pathogens generally grow in high pH environment. A low pH (about 3.5 or less) will prevent the development of some dangerous pathogens in food. For example, in pickled onion, vinegar (pH 2.8) is used. In soft drinks benzoic acid is used for inhibiting the growth of bacteria.
10. **pH in cosmetics:** pH of cosmetics such as vanishing creams, cold creams, skin lotion, skin cleanser, shampoo etc. are carefully maintained for the health of skin and scalp. Any irregularity in the pH of these products will severely damage the skin and scalp.

1.25. BUFFER SOLUTION

A solution whose pH does not change appreciably when small amount of either an acid or a base is added to it is called buffer solution. In other words, solutions that resist the changes in pH are called buffer solutions. Such solutions have reserved pH. For example, $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$, $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$, $\text{CH}_3\text{COONH}_4$ etc.

1.26. TYPES OF BUFFER SOLUTIONS

Buffer solutions are of the following two types:

1. Simple Buffer
2. Mixed Buffer

1.26.1. Simple Buffer

These are the solution of single substances that resist the change in their pH. These are the aqueous solutions of the salts of weak acid and weak base. For example, $\text{CH}_3\text{COONH}_4$, NH_4CN etc.

1.26.2. Mixed Buffer

These are the solutions of the mixtures that resist the change in their pH. Mixed buffer are of two types.

1. Acid buffer
2. Basic buffer

1. Acid Buffer. An acid buffer is a solution of a mixture of a weak acid and a salt of this weak acid with a strong base. For example, $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$, $\text{HCOOH} + \text{HCOOK}$, $\text{H}_2\text{CO}_3 + \text{NaHCO}_3$, Boric acid + Borax (sodium tetraborate), phthalic acid + potassium phthalate etc. The pH of these buffers are less than 7. An equimolar mixture solution of acetic acid and sodium acetate has a pH of 4.76.

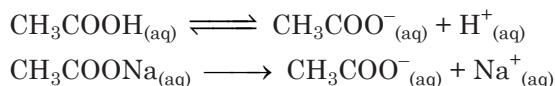
2. Basic Buffer. A basic buffer is a solution of a mixture of weak base and a salt of this weak base with a strong acid. For example, $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$, $\text{NH}_4\text{OH} + \text{NH}_4\text{NO}_3$, Glycine + Glycine hydrochloride etc. The pH of these buffers are more than 7. A mixture of 0.01 M ammonium hydroxide ($K_b = 1.0 \times 10^{-5}$) and 0.1 M ammonium chloride is about 8.

1.27. BUFFER ACTION AND ITS MECHANISM

A buffer solution has to act in a way to remove any hydrogen ions or hydroxide ions that is being added to it, otherwise the pH will change. Acid and basic buffer achieve this in different ways.

1.27.1. Buffer Action of Acid Buffer

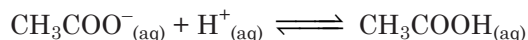
An acid buffer is a solution of a mixture of a weak acid and a salt of this weak acid with a strong base. Let us consider a mixture of CH_3COOH and CH_3COONa that acts as an acid buffer. CH_3COOH and CH_3COONa ionise in its solution as:



Acetic acid is a weak electrolyte and have low value of degree of ionization and therefore the position of the equilibrium will be well to the left. Sodium acetate ionizes to give acetate ion and sodium ion. Due to the common ion effect of CH_3COO^- , the degree of ionization of acetic acid is further suppressed and the position of the equilibrium will shift further to the left according to Le Chatelier's principle. Thus the solution mainly contains CH_3COO^- , Na^+ , undissociated CH_3COOH in excess and few H^+ ions.

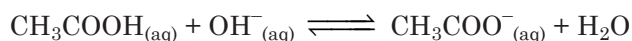
1. Addition of Acid

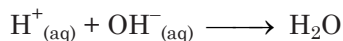
When an acid is added to buffer solution of this type, the H^+ ions furnished by acid react with CH_3COO^- ions to give CH_3COOH which is a weak electrolyte having low value of degree of dissociation. At the same time the dissociation of acetic acid is suppressed due to common ion effect of acetate ion left in the solution. Thus the concentration of H^+ ions in the solution does not change appreciably because most of the new hydrogen ions are removed, and the pH of the solution remains almost unchanged.



2. Addition of alkali

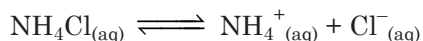
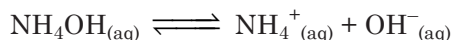
When an alkali is added to buffer solution of this type, the OH^- ions furnished by alkali is removed either by reacting with CH_3COOH to give CH_3COO^- ions and H_2O or by reacting with H^+ ions produced by the ionization of CH_3COOH , and therefore the concentration of OH^- ions does not change appreciably to effect pH.





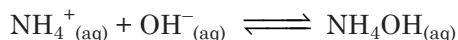
1.27.2. Buffer Action of Basic Buffer

A basic buffer is a solution of a mixture of weak base and a salt of this weak base with a strong acid. Let us consider a mixture of NH_4OH and NH_4Cl that acts as a basic buffer. The NH_4OH and NH_4Cl ionises in its solution as:



Ammonium hydroxide is a weak electrolyte and have low value of degree of ionization and therefore the position of the equilibrium will be well to the left. Ammonium chloride ionizes to give ammonium ion and chloride ion. Due to the common ion effect of NH_4^+ , the degree of ionization of ammonium hydroxide is further suppressed and the position of the equilibrium will shift further to the left according to Le Chatelier's principle.

1. Addition of Alkali. When an alkali is added to buffer solution of this type, the OH^- ions furnished by alkali reacts with NH_4^+ ions to form feebly ionised NH_4OH , whose ionization is further suppressed due to common effect of NH_4^+ ions and therefore the concentration of OH^- ions does not change appreciably to effect pH.



2. Addition of acid

When an acid is added to buffer solution of this type, the H^+ ions furnished by acid is removed either by reacting with NH_4OH to give NH_4^+ ions and undissociated water or by reacting with OH^- ions produced by the ionization of NH_4OH , and therefore the concentration of H^+ ions does not change appreciably to effect pH.

1.28. pH OF BUFFER SOLUTION

The pH of buffer solutions (acid buffer or basic buffer) can be given by an equation called as Henderson-Hasselbalch equation or simply Henderson's equation. The Henderson's equation for acid and basic buffer can be given as below.

$$1. \text{ For acid buffer: } \text{pH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$2. \text{ For basic buffer } \text{pOH} = \text{pK}_b + \log \frac{[\text{salt}]}{[\text{base}]}$$

1.29. CHARACTERISTICS OF BUFFER SOLUTION

1. Buffer solution has definite pH *i.e.*, the acidity or alkalinity of the buffer solution is reserved.
2. The pH of the buffer solution does not change on long standing.
3. The pH of buffer solution is independent of the effect of dilution.
4. The pH of buffer solution does not change appreciably when small amount of either an acid or a base is added to it.
5. Its pH must be in a definite range.

1.30. BUFFER CAPACITY

Buffer solutions are those that resist the changes in pH when either an acid or a base is added to it. Buffer capacity is the measure of the capability of a buffer solution to resist the changes in its pH.

Buffer capacity is quantitatively defined as the number of moles of strong acid or strong base required to change the pH of one litre of a buffer solution by unity. The buffer capacity is represented by ϕ .

$$\text{i.e., } \phi = \frac{\text{Number of moles of strong acid or strong base added to one litre buffer solution}}{\text{Change in pH}}$$

$$\text{or, } \phi = \frac{\Delta B}{\Delta \text{pH}}$$

where, ΔB = No. of moles of acid or base added to one litre buffer solution

ΔpH = Change in pH of the buffer solution

Thus higher the value of buffer capacity, the better is the buffer solution, because a buffer solution with high buffer capacity can accommodate more amount of acid or base without any significant change in pH. Buffer capacity is maximum when:

1. $[\text{Salt}] = [\text{Acid}]$, i.e., $\text{pH} = \text{p}K_a$ for acid buffer
2. $[\text{Salt}] = [\text{Base}]$, i.e., $\text{pOH} = \text{p}K_b$ for base buffer.

Under these conditions the buffer solution is called efficient buffer.

1.31. APPLICATIONS OF BUFFER SOLUTION

Buffer solution find its applications in various fields. Some of the important applications are following:

A. Applications in analytical chemistry

1. In determining pH of solution with the help of indicators.
2. In removing phosphate ion in the qualitative inorganic analysis after second group using $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$ buffer.
3. In precipitating lead chromate quantitatively in gravimetric analysis using buffer solution, $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$.
4. For precipitation of hydroxides of third group in qualitative analysis, a buffer solution, $\text{NH}_4\text{Cl} + \text{NH}_4\text{OH}$, is used.
5. A buffer solution of NH_4Cl , NH_4OH , and $(\text{NH}_4)_2\text{CO}_3$ is used for precipitation of carbonates of fifth group in qualitative inorganic analysis.

B. Applications in pharmaceutical industries

1. Many therapeutic drugs are synthesized under strict pH conditions to ensure the stability and clinical effectiveness of these drugs.
2. To increase the shelf-life of the drugs.
3. In maintaining the drug to near neutral pH so as to avoid the irritation of the tissue.
4. To maintain the drug in a specific unionized or ionized form depending upon the requirement.

C. Applications in other industries

1. In electroplating, the plating of certain alloys can only be achieved successfully if pH is maintained within narrow limits for which buffer solutions are used.
2. Properties of gelatin, which is an important ingredient used in the manufacture of glue, change significantly even with slight change in pH. In order to avoid these changes buffer solutions are used.
3. In printing, buffer solution are used to ensure the ink maintains its normal properties. Inappropriate pH may affect the penetrating and drying properties of the ink.

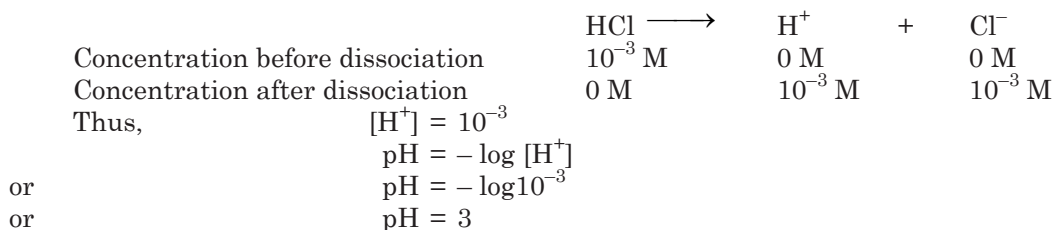
1.32. COROLLARIES

1. $K_w = 1.0 \times 10^{-14}$ mole² litre⁻²
2. $pK_w = 14$
3. $\text{pH} = -\log [\text{H}^+] = -\log [\text{H}_3\text{O}^+]$
4. $\text{pOH} = -\log [\text{OH}^-]$
5. For pure water, $\text{pH} = \text{pOH} = 7$
6. For a solution, $\text{pH} + \text{pOH} = 14$
7. For a solution, $\text{pH} + \text{pOH} = pK_w$
8. For a solution, $[\text{H}^+][\text{OH}^-] = 10^{-14}$ mole² litre⁻²
9. For calculation of pH the concentration must be in moles litre⁻¹
10. The relation between normality and molarity is given as $N = nM$, where, n = valence factor
11. For acids, n = basicity *i.e.*, no. of replaceable H^+ ions
12. For bases, n = acidity, *i.e.*, no. of replaceable OH^- ions
13. If the concentration of H^+ or OH^- ions are less than 10^{-7} mole litre⁻¹ *i.e.*, the concentration of H^+ or OH^- ions in pure water, then for calculation of pH or pOH, total H^+ or OH^- concentration must be used.

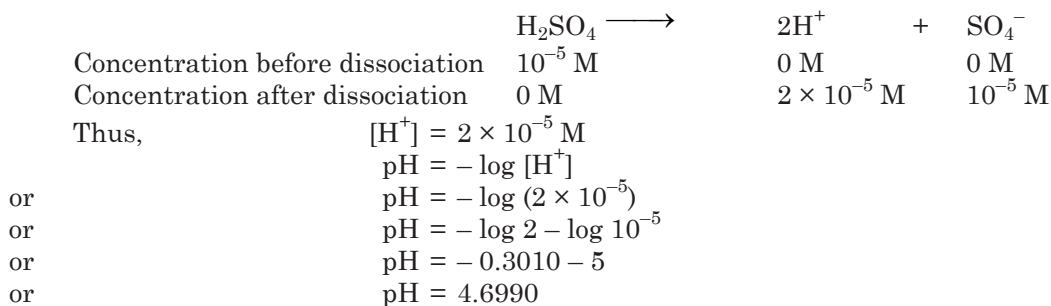
SOLVED EXAMPLES

1. Find the pH of 10^{-3} M HCl

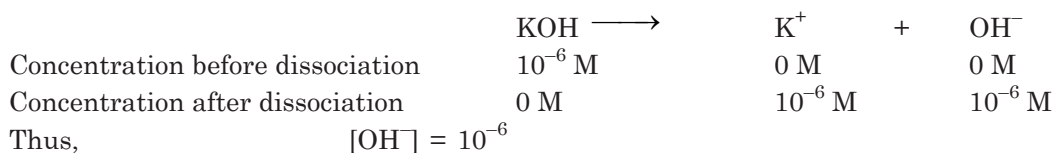
Solution: HCl is a strong electrolyte and ionises completely as below.

2. Find the pH of 10^{-5} M H_2SO_4

Solution: H_2SO_4 is a strong electrolyte and ionises completely as below.

3. Find the pH of 10^{-6} M KOH

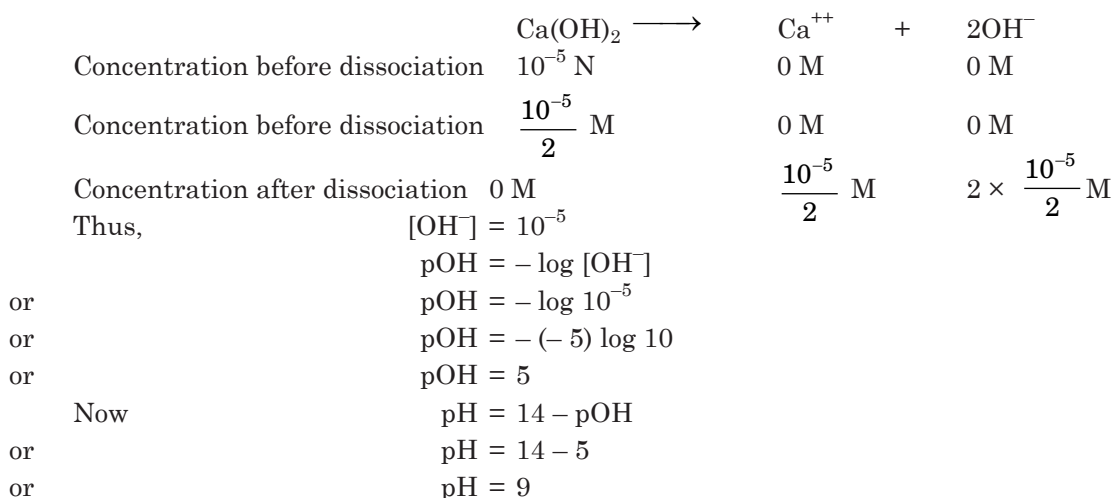
Solution: KOH is a strong electrolyte and ionises completely as below.



$$\begin{aligned} & \text{pOH} = -\log [\text{OH}^-] \\ \text{or} & \text{pOH} = -\log 10^{-6} \\ \text{or} & \text{pOH} = -(-6) \log 10 \\ \text{or} & \text{pOH} = 6 \\ \text{Now,} & \text{pH} = 14 - \text{pOH} \\ \text{or} & \text{pH} = 14 - 6 \\ \text{or} & \text{pH} = 8 \end{aligned}$$

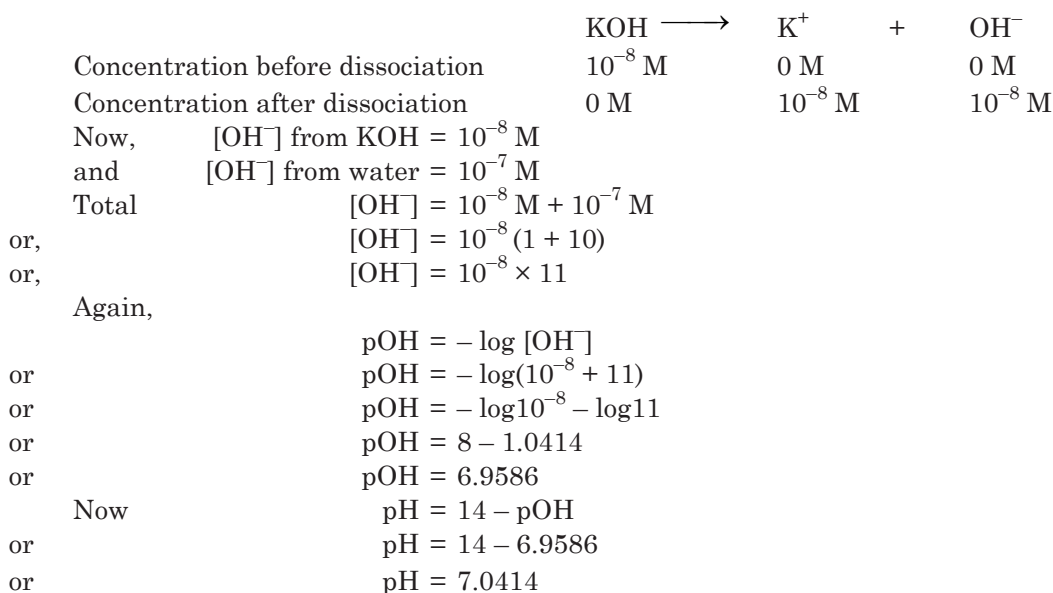
4. Find the pH of 10^{-5} N $\text{Ca}(\text{OH})_2$ assuming complete ionization

Solution: $\text{Ca}(\text{OH})_2$ ionises as below.



5. Find the pH of 10^{-8} M KOH solution

Solution: KOH is a strong electrolyte and ionises completely as below.

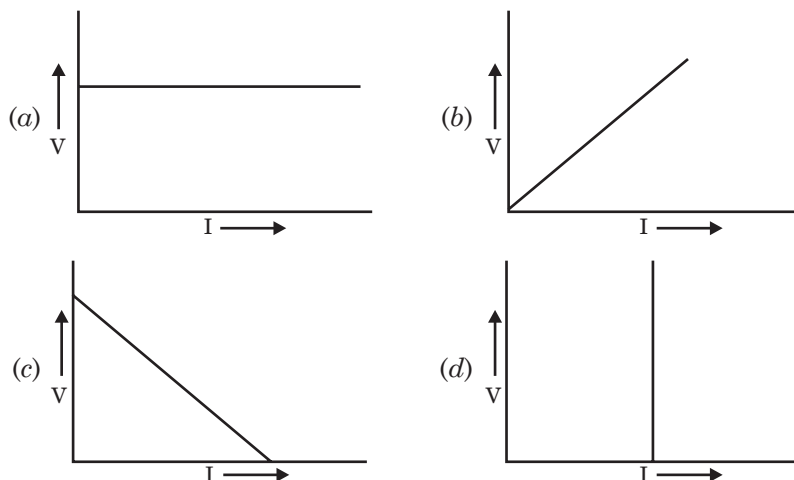


REVIEW EXERCISE

A. Multiple choice question

- Which of the following is an electronic conductor?
(a) aq. NaCl (b) rubber (c) pure water (d) copper
- The non-metal which is electronic conductor:
(a) sulphur (b) plastic (c) glass (d) graphite
- The best known electronic conductor is:
(a) aluminium (b) silver (c) gold (d) copper
- Electronic conductor conducts electricity because of the presence of:
(a) free ions (b) free electrons
(c) both (d) none
- Electrolytic conductors transfer electric current by transfer of:
(a) ions (b) electrons (c) protons (d) all
- An electrolyte is a substance which:
(a) conducts electricity (b) decomposes on heating
(c) is acidic in nature (d) dissociates into ions in aqueous solution
- Theory of ionization was given by
(a) Rutherford (b) Faraday (c) Arrhenius (d) Ostwald
- Electronic conductors transfer electric current by transfer of:
(a) ions (b) electrons (c) protons (d) all
- Dissociation of an electrolyte into its constituent ions in aqueous solution is called:
(a) ionization (b) electrolysis (c) decomposition (d) hydrolysis
- Degree of ionization is equal to:
(a) total number of moles of the electrolyte present in solution
(b) total number of moles of the electrolyte dissociated into ions
(c) number of moles dissociated /total number of moles dissolved
(d) total number of moles dissolved/number of moles dissociated
- The degree of ionization of strong electrolyte is
(a) 1 (b) 0 (c) less than 1 (d) more than 1
- Conductivity of aqueous solution of an electrolyte depends on:
(a) molecular mass of the electrolyte
(b) boiling point of solvent
(c) degree of ionization (d) volume of the solvent
- Degree of ionization does not depend on:
(a) nature of solvent (b) nature of electrolyte
(c) dilution (d) molecular mass of electrolyte
- Substances which give good conducting aqueous solution are called:
(a) weak electrolytes (b) strong electrolytes
(c) non-electrolytes (d) catalysts
- The process in which chemical change occurs on passing electricity is termed:
(a) ionization (b) neutralization
(c) electrolysis (d) hydrolysis
- Strong electrolytes are those which:
(a) dissolve readily in water (b) conduct electricity
(c) dissociate into ions even at high concentration
(d) dissociate into ions at high dilution

17. Which of the following plot correctly represents the Ohm's law:



18. The resistance of a conductor is directly proportional to the:
 (a) length of conductor (b) surface area of conductor
 (c) area of cross-section (d) none
19. As the temperature increase, the resistance of a conductor:
 (a) increases (b) decreases
 (c) sometimes increases sometimes decreases
 (d) none
20. Resistance is the product of specific resistance and:
 (a) conductance (b) cell constant
 (c) specific conductance (d) molar conductance
21. The reciprocal of resistance is called:
 (a) conductance (b) molar conductance
 (c) equivalent conductance (d) specific resistance
22. The unit of conductance is:
 (a) ohm (b) ohm^{-1} (c) ohm^2 (d) ohm^{-2}
23. The resistance offered when the electrodes having area of cross-section 1 cm^2 are apart from each other by 1 cm is called:
 (a) specific resistance (b) cell resistance
 (c) solution resistance (d) specific conductance
24. Unit of specific resistance is:
 (a) $\text{ohm}^{-1} \text{ cm}^{-1}$ (b) ohm cm^{-1}
 (c) ohm cm (d) $\text{ohm}^{-1} \text{ cm}$
25. The reciprocal of specific resistance is called:
 (a) conductance (b) molar conductance
 (c) equivalent conductance (d) specific conductance
26. The conductance offered when the electrodes having area of cross-section 1 cm^2 are apart from each other by 1 cm is called:
 (a) specific conductance (b) molar conductance
 (c) equivalent conductance (d) specific resistance
27. The unit of specific conductance is:
 (a) $\text{ohm}^{-1} \text{ cm}^{-1}$ (b) $\text{ohm}^{-1} \text{ cm}$ (c) ohm cm^{-1} (d) ohm cm

28. The conductance offered by all the ions produced by the ionization of 1 mole of an electrolyte present in a solution is called:
 (a) conductance (b) molar conductance
 (c) equivalent conductance (d) specific conductance
29. The unit of molar conductance is:
 (a) $\text{mho cm}^2 \text{ mole}^{-1}$ (b) $\text{mho cm}^2 \text{ mole}$
 (c) $\text{mho}^{-1} \text{ cm}^2 \text{ mole}^{-1}$ (d) $\text{mho cm}^2 \text{ mole}^{-2}$
30. The conductance offered by all the ions produced by the ionization of 1 gram equivalent of an electrolyte present in a solution is called:
 (a) conductance (b) molar conductance
 (c) equivalent conductance (d) specific conductance
31. The unit of equivalent conductance is:
 (a) $\text{mho cm}^2 \text{ eq}^{-1}$ (b) $\text{mho cm}^2 \text{ eq}$
 (c) $\text{mho}^{-1} \text{ cm}^2 \text{ eq}^{-1}$ (d) $\text{mho cm}^2 \text{ eq}^{-2}$
32. As the concentration of a solution increases its specific conductance:
 (a) decreases (b) increases
 (c) sometimes increases sometimes decreases
 (d) none
33. The molar and equivalent conductance of strong electrolytes:
 (a) increases only slightly with dilution
 (b) increases appreciably with dilution
 (c) decreases slightly with dilution
 (d) decreases appreciably with dilution
34. The molar and equivalent conductance of weak electrolytes
 (a) increases largely with dilution
 (b) decreases largely with dilution
 (c) increases slightly with dilution
 (d) decreases slightly with dilution
35. The product of the concentrations of H^+ and OH^- of water is called
 (a) ionic product of water (b) solubility product
 (c) concentration product (d) none
36. The ionic product of water at 298K is:
 (a) $1.0 \times 10^{-14} \text{ mole}^2 \text{ litre}^{-2}$ (b) $1.0 \times 10^{-7} \text{ mole}^2 \text{ litre}^{-2}$
 (c) $7.0 \times 10^{-14} \text{ mole}^2 \text{ litre}^{-2}$ (d) $14.0 \times 10^{-7} \text{ mole}^2 \text{ litre}^{-2}$
37. The concentration of H^+ in pure water is:
 (a) $10^{-14} \text{ mole litre}^{-1}$ (b) $10^{-14} \text{ mole}^2 \text{ litre}^{-2}$
 (c) $10^{-7} \text{ mole litre}^{-1}$ (d) $10^{-7} \text{ mole}^2 \text{ litre}^{-1}$
38. Sum of pH and pOH of pure water is:
 (a) 14 (b) 9 (c) 5 (d) 7
39. The pOH of pure water:
 (a) 7 (b) 10 (c) 14 (d) 3
40. The value of pK_w for pure water is:
 (a) $10^7 \text{ mole}^2 \text{ litre}^{-2}$ (b) $10^{-9} \text{ mole}^2 \text{ litre}^{-2}$
 (c) $10^{-14} \text{ mole}^2 \text{ litre}^{-2}$ (d) $10^0 \text{ mole}^2 \text{ litre}^{-2}$
41. Solution of fairly constant pH is called:
 (a) salt solution (b) acidic solution
 (c) buffer solution (d) basic solution

42. A solution that resist to change of its pH value:
(a) dil. HCl (b) dil. NaOH
(c) buffer solution (d) dil. H₂SO₄
43. If pH of any solutions is equal to zero, then solution will be
(a) acidic (b) basic
(c) neutral (d) none of these
44. If the K_w of boiling water is 10^{-12} mole² litre⁻², then its pH will be:
(a) 12 (b) 2 (c) 8 (d) 6
45. pH of water at 25°C is:
(a) 12 (b) 2 (c) 7 (d) 10
46. The pH of a solution can be defined as:
(a) negative logarithm of H⁺ ion concentration
(b) positive logarithm of H⁺ ion concentration
(c) H⁺ ion concentration
(d) negative logarithm of 1/H⁺ ion concentration
47. The pH of a solution is 6, its pOH will be:
(a) 6 (b) 8 (c) 14 (d) 0
48. A solution having equal pH and pOH is called:
(a) dilute solution (b) buffer solution
(c) neutral solution (d) isotonic solution
49. The value of pK_w is 14.6. If the pH of a solution is 7, then the nature of solution will be:
(a) acidic (b) basic
(c) neutral (d) can't be predicted
50. The strength of an acid depends upon:
(a) concentration of acid (b) concentration of H⁺ ion in solution
(c) hydrolysis
(d) number of moles of base used in the neutralisation.
51. A solution will be neutral if:
(a) H⁺ ions are not present (b) OH⁻ ions are not present
(c) both H⁺ ions OH⁻ ions are not present
(d) both H⁺ ions OH⁻ ions are in equal concentration
52. The value of H⁺ ion concentration for a neutral solution is:
(a) 1×10^{-7} mole litre⁻¹ (b) 1×10^{-14} mole litre⁻¹
(c) 2×10^{-14} mole litre⁻¹ (d) 0.5×10^{-7} mole litre⁻¹
53. If the ionic product of water (K_w) be 1.5×10^{-14} mole² litre⁻², then the pH of water will be:
(a) between 6 to 7 (b) between 7 to 8
(c) 7 only (d) 8 only
54. At 25°C the pH of water is 7. If water is heated to 50°C its:
(a) pH will decrease (b) pH will increase
(c) [H⁺] will be more than [OH⁻] (d) [H⁺] will be more than [OH⁻]
55. When temperature increases, ionic product of water:
(a) will decrease (b) will increase
(c) will remain same (d) none of these
56. The ionic product of water is expressed in
(a) mole⁻¹ litre⁻¹ (b) mole⁻² litre⁻²
(c) mole² litre⁻² (d) mole litre⁻¹

57. If the temperature of pure water increases
 (a) pH increases pOH increases (b) pH decreases pOH decreases
 (c) pH increases pOH decreases (d) pH decreases pOH increases
58. An aqueous solution with pH 2.2 will be:
 (a) acidic (b) alkaline
 (c) neutral (d) amphoteric
59. The ionic product of water in an aqueous solution of NaCl:
 (a) more than 10^{-14} (b) less than 10^{-14}
 (c) 10^{-14} (d) zero
60. Which one of the following will be alkaline:
 (a) solution having $[H^+] = 10^{-7}$ (b) solution having $[H^+] = 10^{-12}$
 (c) solution having $[H^+] = 10^{-4}$ (d) solution having $[H^+] = 10^{-6}$
61. A solution of pOH 12 is acidic than a solution of pOH 8
 (a) more (b) less (c) equally (d) none
62. A buffer solution of NH_4Cl and NH_4OH is diluted with water, its pH
 (a) decrease (b) increase (c) remains same (d) becomes 7
63. If small amount of either an acid or a base is added to an unknown solution, its pH remains unaltered, the solution is a:
 (a) isotonic solution (b) dilute solution
 (c) isomorphous solution (d) buffer solution
64. Which of the following combination does not form a buffer solution:
 (a) $CH_3COOH + CH_3COONa$ (b) $NH_4OH + NH_4Cl$
 (c) $NaOH + NaCl$ (d) $NH_4OH + (NH_4)_2SO_4$
65. Highest pH value stands for
 (a) strongly acidic solution (b) strongly alkaline solution
 (c) neutral solution (d) weakly acidic solution

B. Answer the following Questions

- Define conductor and insulator with example.
- Explain different types of conductor with suitable example.
- Distinguish between electronic and electrolytic conductor.
- Explain electrolytes and non-electrolytes.
- What do you mean by strong and weak electrolytes? Explain with example.
- Write down the important features of Arrhenius theory of electrolytic dissociation.
- State and explain Ohm's law.
- What do you mean by resistance? Explain various factors affecting resistance of a conductor.
- What do you mean by cell constant? What is its significance?
- Explain the term conductance.
- Explain specific resistance and specific conductance. Write down their units also.
- What do you mean by molar conductance and equivalent conductance? Explain.
- How specific conductance changes with solution? Explain.
- Explain the effect of dilution on molar conductance and equivalent conductance.
- What do you mean by ionic product of water? Explain.
- Define pH. Explain the nature of solution on the basis of pH.
- What do you mean by pOH? How the nature of a solution is related with pOH?
- What is meant by pH scale? Write down the limitations of pH scale.
- Write down some important applications of pH.
- What is buffer solution? What are their different types? Explain.

21. Explain acid buffer and basic buffer with suitable example.
22. Explain the mechanism of buffer action.
23. What do you understand by buffer capacity? Explain.
24. Write down the important characteristics of buffer solution.
25. Write down some important applications of buffer solution.

C. Numerical Questions

1. Calculate the pH of the following solutions assuming complete ionization.

(a) 10^{-5} M HCl	(b) 10^{-3} N HNO ₃
(c) 10^{-6} N H ₂ SO ₄	(d) 10^{-4} M H ₃ PO ₄
2. Calculate the pH of the following solutions assuming complete ionization.

(a) 10^{-5} M NaOH	(b) 10^{-6} N KOH
(c) 10^{-5} N Ca(OH) ₂	(d) 10^{-4} M Ba(OH) ₂
3. Calculate the pH of the following solution

(a) 10^{-8} M HCl	(b) 10^{-8} M NaOH
(c) 10^{-9} M HNO ₃	(d) 10^{-9} M KOH
4. Calculate the pH of the following solutions.

(a) 0.15 M HCl	(b) 0.15 M H ₂ SO ₄
(c) 2.5×10^{-5} M NaOH	(d) 2.8 gram KOH in 500 ml solution
5. Calculate the pH of 0.001 N NaOH solution in water
6. Calculate the pH of 0.0001 N HCl solution in water
7. The pH of a solution is 10. What is the concentration of H⁺ ions in the solution?
8. The pH of a solutions is 5. Find the concentration of H⁺ and OH⁻ ions in the solution.

Answers**A. Multiple Choice Questions**

1	2	3	4	5	6	7	8	9	10	11	12
(d)	(d)	(b)	(b)	(a)	(d)	(c)	(b)	(a)	(c)	(a)	(c)
13	14	15	16	17	18	19	20	21	22	23	24
(d)	(b)	(c)	(c)	(b)	(a)	(a)	(b)	(a)	(b)	(a)	(c)
25	26	27	28	29	30	31	32	33	34	35	36
(d)	(a)	(a)	(b)	(a)	(c)	(a)	(b)	(a)	(a)	(a)	(a)
37	38	39	40	41	42	43	44	45	46	47	48
(c)	(a)	(a)	(c)	(c)	(c)	(a)	(d)	(c)	(a)	(b)	(c)
49	50	51	52	53	54	55	56	57	58	59	60
(a)	(b)	(d)	(a)	(a)	(a)	(b)	(c)	(b)	(a)	(c)	(b)
61	62	63	64	65							
(a)	(c)	(d)	(c)	(a)							

C. Numerical Questions

1. (a) 5 (b) 3 (c) 6 (d) 3.52
2. (a) 9 (b) 8 (c) 9 (d) 3.69
3. (a) 6.96 (b) 7.04 (c) 6.996 (d) 7.004
4. (a) 0.82 (b) 0.552 (c) 9.398 (d) 13
5. 11 6. 4 7. 10^{-10} mole lt⁻¹
8. 10^{-5} mole lt⁻¹, 10^{-9} mole lt⁻¹

A Textbook of **ENGINEERING CHEMISTRY-II**

As per the Latest Syllabus of Diploma in Engineering Courses Under
Jharkhand University of Technology, Ranchi

About the Book:

The objective of the present book is to serve the students with an elementary knowledge of chemistry. The syllabi of chemistry taught in the name of engineering chemistry in different diploma colleges are fairly diverse in nature. It is therefore quite impossible to provide a complete coverage at all the topics in a limited space for all students. However, author has dealt with modern views of the topics of the syllabi and attempted to give a major coverage at the recent syllabi taught in various diploma level institutions.

About the Author:



Dr. Altaf Ahmad obtained M.Sc. degree in Chemistry (Special-Organic Chemistry) from Ranchi University, Ranchi and M.Tech. degree in Surface Engineering from NIT, Jamshedpur. He completed his Ph.D. (Corrosion Science and Engineering) from Ranchi University, Ranchi, Jharkhand. Dr. Altaf is a Senior Lecturer and Incharge of Department of Chemistry, Al-Kabir Polytechnic, Jamshedpur, Jharkhand. He has been teaching Chemistry for last 25 years. He is also associated with research work and published many research papers in the Journals of

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ISO 9001:2015

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