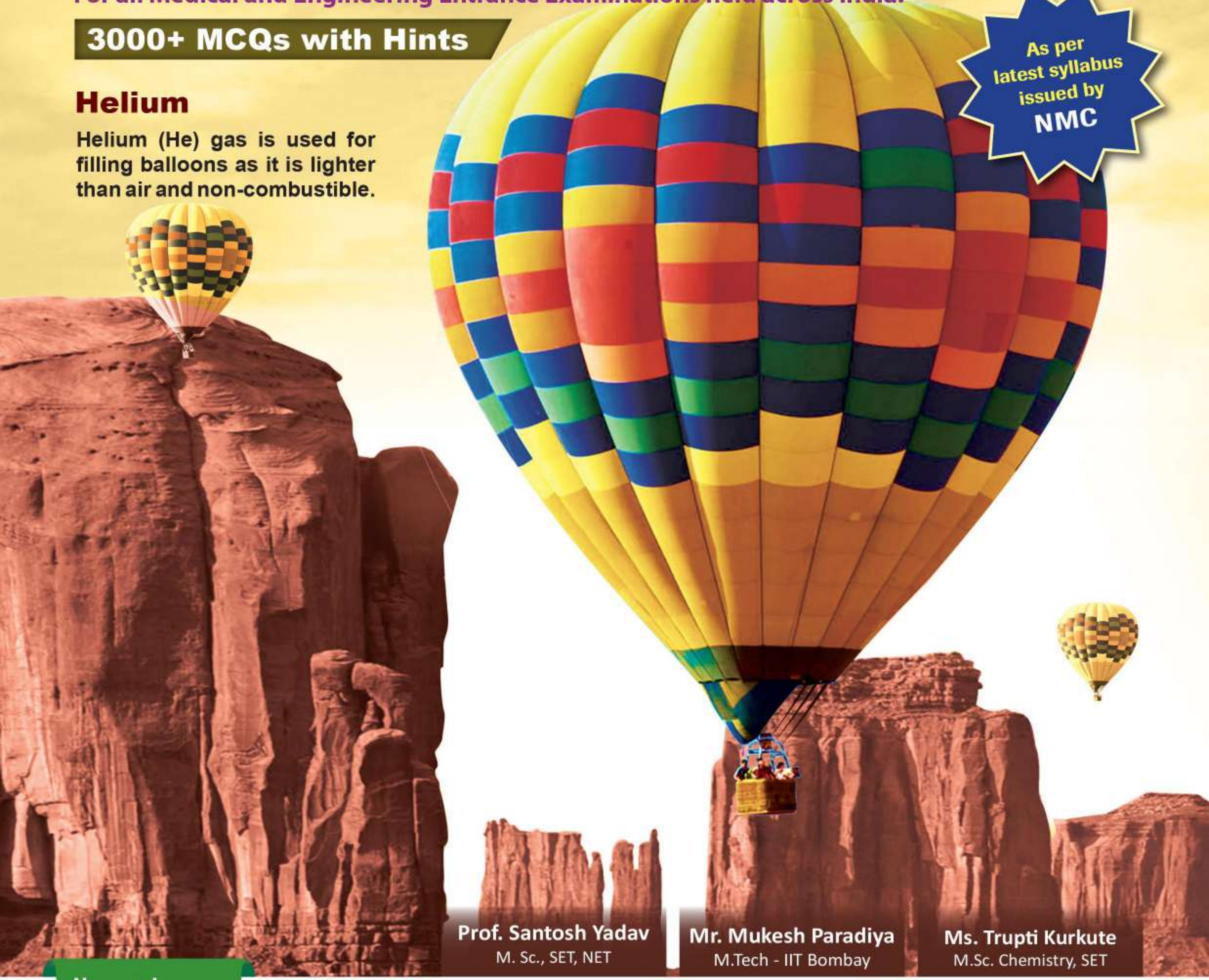
NEET-UG & JEE (Main)
CHEMISTRY Vol - II

For all Medical and Engineering Entrance Examinations held across India.

3000+ MCQs with Hints

As per latest syllabus issued by NMC



Helium

Helium (He) gas is used for filling balloons as it is lighter than air and non-combustible.

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Now with more study techniques

Target Publications® Pvt. Ltd.

For all Medical and Engineering Entrance Examinations held across India.

Absolute
NEET – UG & JEE (Main)
Chemistry Vol. II

Now with more study techniques

Updated as per latest syllabus for:
NEET (UG) 2024 issued by **NMC on 6th October, 2023**
JEE (Main) 2024 issued by **NTA on 1st November, 2023**

Salient Features

- Comprehensive theory for every topic
- Subtopic-wise segregation of MCQs for efficient practice
- Exhaustive coverage of questions from previous NEET (UG), JEE (Main) and other competitive examinations **till year 2024**:
  - **3223** MCQs
  - **83** Numerical Value Type (NVT) questions
  - Solutions to the questions are provided for better understanding
- Includes **Smart Keys:** Multiple study techniques to enhance understanding and problem solving:
  - ***Smart code***
  - ***Smart tip***
  - ***Caution***
  - ***Remember This***
  - ***Think out of the box***
- Includes relevant Solved Questions from:
  - NEET (UG) 2023
  - JEE (Main) 2023 24th Jan (Shift - II)
  - NEET (UG) Manipur 2023
- Includes Question Paper and Answer Keys (Solutions through Q.R. code) of:
  - JEE (Main) 2024 31st January (Shift - I)
  - NEET (UG) 2024
- Topic Test provided in each chapter for self-assessment
- Q.R. codes provide:
  - Video links for boosting conceptual retention
  - Answers & Solutions to Topic Tests

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PREFACE

Target's **'Absolute Chemistry Vol - II'** is a complete guidebook, extremely handy for preparation of NEET (UG) and JEE (Main) examinations. This edition provides an unmatched comprehensive amalgamation of theory with MCQs. The chapters are aligned with the latest syllabus for **NEET (UG) and JEE (Main) 2024** examinations. Although the alignment runs parallel to NCERT curriculum, the structure of the chapters prioritizes knowledge building of the students. The book provides the students with scientifically accurate context, several study techniques and skills required to excel in these examinations.

All the questions included in a chapter have been specially created and compiled to enable students solve complex problems which require strenuous effort with promptness.

These MCQs are framed considering the importance given to every topic as per the NEET-UG & JEE (Main) exam to form a strong foundation. They are a healthy mix of theoretical, numerical, reactions and graphical based questions.

The level of difficulty of these questions is at par with that of various competitive examinations held across India. Questions from various examinations such as NEET (UG), JEE (Main), MHT CET, KCET, WB JEE, AP EAMCET, TS EAMCET, AP EAPCET, GUJ CET are exclusively covered.

Features in each chapter:

- Coverage of **'Theoretical Concepts'** that form a vital part of any competitive examination.
- **'Quick Review'** which highlights the key concepts of the chapter in the form of tables/ flow charts aids in last-minute revision.
- **'Formulae'** covers all the key formulae in the chapter, making it useful for students to glance at while solving problems.
- **'Multiple Choice Questions'** are segregated topic-wise to enable easy assimilation of questions based on the specific concept.
- **'Topic Test'** has been provided at the end of each chapter to assess the level of preparation of the student on a competitive level.

All the features of this book pave the path of a student to excel in examination. The features are designed keeping the following elements in mind: Time management, easy memorization or revision and non-conventional yet simple methods for MCQ solving.

We hope the book benefits the learner as we have envisioned.

A book affects eternity; one can never tell where its influence stops.

Publisher

Edition: Eighth

The journey to create a complete book is strewn with triumphs, failures and near misses. If you think we've nearly missed something or want to applaud us for our triumphs, we'd love to hear from you.

Please write to us on: mail@targetpublications.org

Disclaimer

This reference book is based on the NEET (UG) and JEE (Main) syllabus prescribed by National Testing Agency (NTA). We the publishers are making this reference book which constitutes as fair use of textual contents which are transformed by adding and elaborating, with a view to simplify the same to enable the students to understand, memorize and reproduce the same in examinations.

This work is purely inspired upon the course work as prescribed by the National Council of Educational Research and Training (NCERT). Every care has been taken in the publication of this reference book by the Authors while creating the contents. The Authors and the Publishers shall not be responsible for any loss or damages caused to any person on account of errors or omissions which might have crept in or disagreement of any third party on the point of view expressed in the reference book.

KEY FEATURES

Quick Review

'**Quick Review**' Quick review includes tables/ flow chart to summarize the key points in chapter.

Quick Review

Formulae

Formulae

'**Formulae**' include all of the important formulae in the chapter, which helps students quickly and easily solve problems and revise at the last minute.

Examples

'**Examples**' are provided to elucidate a concept using theory or MCQ or numerical question and help students hone their problem-solving skills.

Examples

Smart Keys

Smart Keys

Smart Keys comprise a set of remarkable study techniques contrived to benefit students.

Caution

'**Caution**' apprises students about mistakes which are made while solving an MCQs.

Caution

Smart Tips

Smart Tips

'**Smart Tips**' comprise important theoretical or formula based short tricks considering their usage in solving MCQ.

Think out of the box

'**Think out of the box**' reveals quick witted approach to crack the specific question.

Think out of the box

Smart Code

Smart Code

'**Smart Code**' showcases simple and smart mnemonic created for selected concepts.

Remember This

'**Remember This**' includes key points, important reactivity orders, exceptions, point of difference, misconceptions, etc.

Remember This

To be continued…

For your knowledge

For your Knowledge

'**For your Knowledge**' includes additional information relevant to concept.

Connection

Connection

'**Connection**' enables students to interlink concepts covered in different chapters.

Gyan Guru

Gyan Guru

'**Gyan Guru**' illustrates real life applications or examples related to the concept discussed.

Clock

Clock symbol

'**Clock Symbol**' instructs students that given MCQ can be solved apace by applying either smart tips or think of the box.

QR Code

QR Code

'**QR Code**' provides:

i. Access to a video/PDF in order to boost understanding of a concept or activity.

ii. Answers & Solutions to Topic Test of each chapter.

iii. Solutions to question papers of

a. JEE (Main) 2024 31st January (Shift - 1)

b. NEET (UG) 2024

Frequently Asked Questions

- **Why Absolute Series?**

  Gradually, every year the nature of competitive entrance exams is inching towards conceptual understanding of topics. Moreover, it is time to bid adieu to the stereotypical approach of solving a problem using a single conventional method.

  To be able to successfully crack the NEET/JEE (Main) examinations, it is imperative to develop skills such as data interpretation, appropriate time management, knowing various methods to solve a problem, etc. With Absolute Series, we are sure, you'd develop all the aforementioned skills and take a more holistic approach towards problem solving. The way you'd tackle advanced level MCQs with the help of Hints, Solved examples, Smart tips, Smart codes and Think out of the box would give you the necessary practice that would be a game changer in your preparation for the competitive entrance examinations.

- **What is the intention behind the launch of Absolute Series?**

  The sole objective behind the introduction of Absolute Series is to cater to needs of students across a varied background and effectively assist them to successfully crack the NEET/JEE (Main) examinations. With a healthy mix of MCQs, we intend to develop a student's MCQ solving skills within a stipulated time period.

- **What do I gain out of Absolute Series?**

  After using Absolute Series, students would be able to:

  a. assimilate the given data and apply relevant concepts with utmost ease.
  b. tackle MCQs of different pattern such as match the columns, diagram based questions, multiple concepts and assertion-reason efficiently.
  c. garner the much needed confidence to appear for competitive exams.
  d. easy and time saving methods to tackle tricky questions will help ensure that time consuming questions do not occupy more time than you can allot per question.

- **How to derive the best advantage of the book?**

  To get the maximum benefit of the book, we recommend :

  a. Go through the detailed theory and Examples solved alongwith at the beginning of a chapter for concept clarity. Commit Smart Tips into memory and pay attention to Caution, Remember This.
  b. Read through the Quick review section to summarize the key points in chapter.
  c. Know all the Formulae compiled at the end of theory by heart.
  d. Using subtopic wise segregation as a leverage, work through MCQs in each subtopic by applying the concepts of chemistry. Questions from exams such as JEE (Main), NEET-UG are tagged and placed along the flow of subtopic. Mark these questions specially to gauge the trends of questions in various exams.
  e. Be extra receptive to Alternate Method and application of Smart Keys, assimilate them into your thinking.

Best of luck to all the aspirants!

2 Electrochemistry

INTRODUCTION

➢ **Electrochemistry:**

> ***Electrochemistry*** *is the branch of chemistry which deals with the study of production of electrical energy (electricity) from chemical energy produced by spontaneous chemical reactions and use of the electrical energy to bring about non-spontaneous chemical reactions.*

➢ **Importance of electrochemistry:**

i. Study of electrochemistry is a vast and interdisciplinary subject of immense importance not only for theoretical studies but also for practical purposes.

ii. A large number of chemicals are produced by this method.

E.g. a. Extraction of metals like sodium, potassium, aluminium, copper is carried out by electrochemical method.

b. Extraction of non-metals like chlorine, fluorine, etc.

c. Production of chemicals like sodium hydroxide.

iii. Electrochemistry also finds its application in purification (electro-refining) of metals like copper, silver, gold, aluminium, etc. and in electroplating wherein one metal is coated on the surface of another metal.

iv. Different types of batteries are used for different purposes that produce electrical energy from chemical energy.

v. Electrochemical reactions can be energy efficient and less polluting. Therefore, study of electrochemistry has attracted focus for creating new ecofriendly technologies.

vi. Fuel cells are used as high efficiency and less polluting alternate source of electrical power.

vii. Communication and transmission of sensory signals at cellular level in organisms is also of electrochemical origin.

REDOX REACTIONS

➢ **Redox reactions:**

i. The reactions, which are brought about by loss of electrons (oxidation) and gain of electrons (reduction) simultaneously, are called **oxidation-reduction reactions** or **redox reactions.**

OR

The reactions that involve change in oxidation number of the interacting species are known as **redox reactions.**

ii. These redox reactions are made up of two half reactions; one involving the loss of electrons, known as oxidation half reaction and other involving the gain of electrons known as reduction half reaction. Electrons are transferred from one reactant to another.

CONTENTS

iii. During oxidation, the oxidation number of an element increases due to loss of electrons. During reduction, the oxidation number of an element decreases due to gain of electrons.

iv. In these reactions, the substance which loses one or more electrons (i.e., gets oxidized), acts as a reducing agent by supplying electrons and is known as a **reductant**.

v. However, the substance which gains one or more electrons (i.e., gets reduced), acts as an oxidizing agent by accepting electrons and is known as an **oxidant**.

E.g. $2Na + Cl_2 \longrightarrow 2Na^+ + 2Cl^-$

The above reaction is composed of two half reactions:

Oxidation Half Reaction

$2Na \longrightarrow 2Na^+ + 2e^-$ **(oxidation)**

$\underbrace{\overset{0}{} \quad \text{loss of } e^- \quad \overset{+1}{}}$

(Oxidation number increases)

Reducing agent – Na (by supplying electrons)

Reduction Half Reaction

$Cl_2 + 2e^- \longrightarrow 2Cl^-$ **(reduction)**

$\underbrace{\overset{0}{} \quad \text{gain of } e^- \quad \overset{-1}{}}$

(Oxidation number decreases)

Oxidizing agent – Cl_2 (by accepting electrons)

Connections

In Chapter 7, Redox Reactions (Vol. I), you have learnt about the basics and balancing of redox reactions.

ELECTROCHEMICAL CELLS

- **Electrochemical cells:**

  Electrochemical reactions (redox reactions) occurring in an electrochemical cell involve the transfer of electrons from one species to the other.

  There are two types of electrochemical cells:

Electrochemical cells

Galvanic or Voltaic cells	Electrolytic cells
• Conversion of chemical energy to electrical energy. • Positive electrode – Cathode – Reduction Negative electrode – Anode – Oxidation • Spontaneous chemical reaction occurs by itself and generates electricity. • **E.g.** Dry cell, fuel cell, lead storage cell, etc.	• Conversion of electrical energy to chemical energy • Positive electrode – Anode – Oxidation Negative electrode – Cathode – Reduction • Non-spontaneous reaction is forced to occur by electric current from an external source. • **E.g.** Electrorefining, electroplating, etc.

Terms	Description
Current	Flow of electrons through a wire or any conductor.
Electrode	A metallic strip that provides surface for oxidation or reduction reaction or simply acts as electron conductor without undergoing chemical reaction (inert electrode.)
Anode	Oxidation occurs at anode.
Cathode	Reduction occurs at cathode.
Electrolyte	Salt solution in which electrodes are dipped.
Half-cell	Each half-cell consists of an electrode dipped into an electrolyte. Each electrochemical cell consists of two half-cells.
Salt bridge	Connects the two half-cells.

Smart code - 1

Alphabetically **Anode** precedes **Cathode** and **Oxidation** precedes **Reduction**.

Anode is where oxidation occurs and cathode is where reduction occurs.

GALVANIC OR VOLTAIC CELLS

➢ **Galvanic/Voltaic cells:**

i. In galvanic cells, electrical current is generated by spontaneous redox reactions.

ii. **Daniel cell** is a type of Galvanic cell, in which, the spontaneous reaction of zinc metal with an aqueous solution of copper sulphate is used.

$$Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

iii. The Daniel cell consists of two half-cells:
Reduction half-cell: Copper rod (copper electrode) dipped in 1 M copper sulphate solution, and
Oxidation half-cell: Zinc rod (zinc electrode) dipped in 1 M zinc sulphate solution.

iv. The two half-cells are connected by a metallic wire through a voltmeter and a switch externally.

v. The electrolytes of the two half cells are connected internally through a salt bridge as shown in the figure.

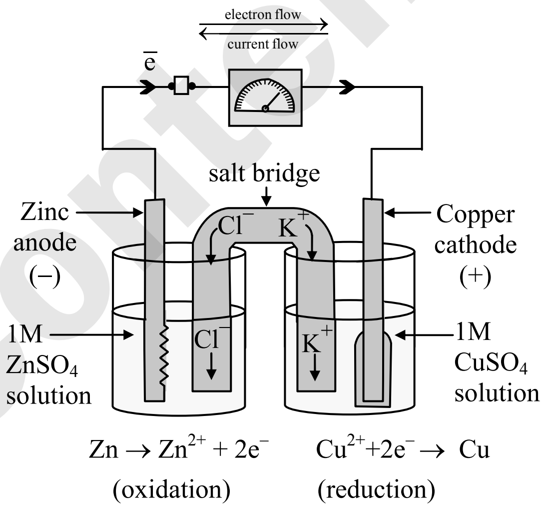
vi. During the reaction, zinc is oxidized to Zn^{2+} ions which go into the solution.
Therefore, $Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^-$ (Oxidation half reaction)

vii. The electrons released at the zinc electrode move towards the other electrode through external circuit. These electrons are accepted by Cu^{2+} ions of $CuSO_4$ and the ions get reduced to metallic Cu.
$Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$ (Reduction half reaction)

viii. The electrode at which oxidation occurs is called anode; it becomes negatively charged due to released electrons.

ix. However, the electrode at which reduction occurs is called cathode; it becomes positively charged as electrons are consumed.

x. When the complete cell is set up and the switch is in 'on' position, flow of electrons takes place from zinc electrode to copper electrode in the external circuit. Due to this, zinc dissolves in oxidation half-cell solution to form Zn^{2+} ions. In reduction half-cell, Cu^{2+} ions pick up the electrons, get converted to metallic copper and deposit on the cathode.



A simple galvanic cell (Daniel cell)

xi. The electrical potential of Daniel cell is **1.1 V**.

xii. Application of an external opposite potential to the galvanic cell with a slow increase in its value shows that the reaction continues to take place until the opposing voltage reaches the value 1.1 V. At this potential, the reaction stops altogether and no current flows through the cell. Any further increase in the external potential again starts the reaction but in the opposite direction and it now functions as an **electrolytic cell,** a device which uses electrical energy to bring about non-spontaneous chemical reactions.

xiii. The salt bridge maintains electrical neutrality in both the compartments by flow of ions. Two Cl^- ions from the salt bridge migrate into the anode solution for every Zn^{2+} ion formed and two K^+ ions from the salt bridge migrate into the cathode solution to replace every Cu^{2+} ion reduced. Neither K^+ nor Cl^- ions are reduced or oxidized in preference to the Cu^{2+} ions or Zn atoms. This flow of ions completes the electrical circuit and ensures the complete supply of current.

xiv. A large number of galvanic cells can be constructed in this way by taking combinations of different half-cells. The electrolytes of two half-cells may be connected through a salt bridge (**E.g.**, Daniel cell) or sometimes both the electrodes are dipped in the same electrolytic solution wherein no salt bridge is required (**E.g.**, fuel cell).

➢ **Salt bridge:**

i. Salt bridge is a U-shaped glass tube which connects the two half-cells of a galvanic cell.

ii. The glass tube is filled with saturated solution of electrolyte like KCl, KNO_3 or NH_4NO_3, etc., (whose ions have almost same mobility i.e., transport number) prepared in agar-agar jelly or gelatin.

iii. The two openings of U-tube are plugged with some porous material such as glass wool or cotton.

iv. Functions of salt bridge:
a. Salt bridge completes the electrical circuit by providing ionic contact between the two solutions.
b. It prevents the mixing of the electrolytic solutions.
c. It maintains electrical neutrality of the two half-cell solutions by flow of ions.

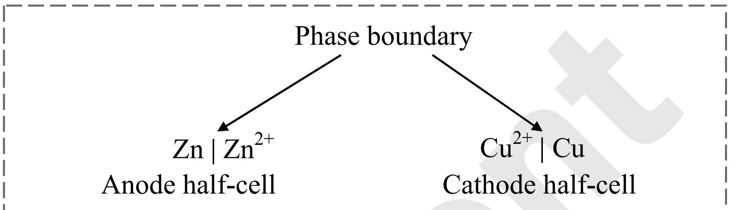
v. If the salt bridge is not connected and two half-cells are connected with metal wire, then electrons produced in anode compartment flow to the cathode compartment through a wire. The current flows for an instant and then stops. This is because the electrons leave the left compartment that becomes positively charged and the right compartment becomes negatively charged by receiving electrons. Due to the accumulation of charge in two compartments, the flow of electrons (current) eventually stops.

➢ **Cell notation or representation of a galvanic cell:**

According to IUPAC conventions, a galvanic cell can be represented symbolically as follows:

i. Each half-cell is represented by putting a vertical line between the metal electrode and the electrolytic solution. This vertical line represents the phase boundary. **E.g.** $Zn \mid Zn^{2+}$

It is important to note that metal electrode in anode half-cell is written on the left of metal ion, while in cathode half-cell, it is written on the right of metal ion.



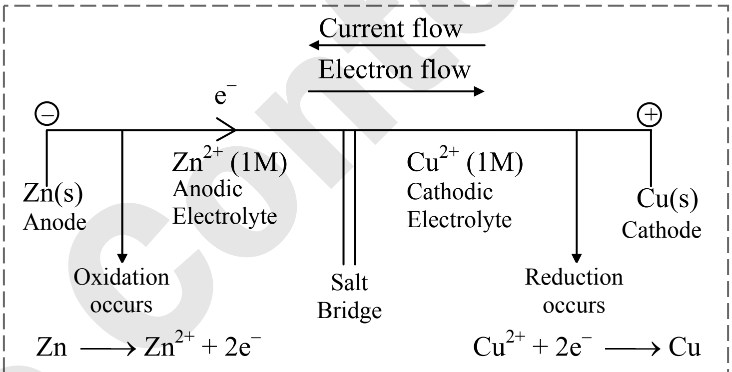
Representation of Anode and Cathode Half-cells

ii. Now, for representing a cell, **anode half-cell** is written on the **left** while **cathode half-cell** is written on **right**.

iii. In a complete cell diagram, the two half-cells are separated by a double vertical line (representing the salt bridge) in between.

iv. Therefore, zinc–copper cell can be represented as:
$Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$.

v. For more information, the physical state of the different species involved are mentioned. The concentration of the electrolyte is also mentioned within bracket after the cation, i.e., $Zn(s) \mid Zn^{2+}(1\ M) \parallel Cu^{2+}(1\ M) \mid Cu(s)$



Complete cell representation

vi. The cell reaction for the above cell notation is $Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$

vii. For gaseous reactions, an inert electrode such as platinum is used, then the hall-cell is represented as:
$Pt(s) \mid H_2(g) \mid H^+(aq)$

➢ **Types of electrodes:**

Electrodes

- **Metal–metal ions electrode**
  - A metal strip dipped in a solution of metal ions. **E.g.** $Zn^{2+}(aq) \mid Zn(s)$
- **Metal-sparingly soluble salt electrode**
  - Metal coated with its sparingly soluble salt dipped in a solution of soluble salt having common anion with that of sparingly soluble salt.

  E.g. Silver-silver chloride electrode: $Cl^-(aq) \mid AgCl(s) \mid Ag(s)$
- **Gas electrode (non metal-non metal ion electrode)**
  - A gas bubbled through an inert metal electrode which is immersed in a solution having same cation as that of gas.

  E.g. Hydrogen gas electrode:
  $H^+(aq) \mid H_2(g, P_{H_2}) \mid Pt(s)$ (Here, P_{H_2} represents pressure of hydrogen gas)
- **Redox electrode**
  - A platinum wire dipped in solution containing the ions of the same substance in two different oxidation states. **E.g.** $Cu^{+}(aq),\ Cu^{2+}(aq) \mid Pt(s)$

  Note: The order of writing the two ions is immaterial.

ELECTRODE POTENTIAL AND CELL POTENTIAL

Electrode potential:

The tendency of an electrode to lose or gain electrons when it is in contact with its own ions in solution is called ***electrode potential****.*

OR

The potential difference that develops between electrode and electrolyte during oxidation and reduction reaction at electrode is called ***electrode potential****.*

The tendency of an electrode to lose electrons or to get oxidized is called its **oxidation potential**.

$M_{(s)} \rightleftharpoons M^{n+}_{(aq)} + ne^-$

The tendency of an electrode to gain electrons or to get reduced is called its **reduction potential**.

$M^{n+}_{(aq)} + ne^- \rightleftharpoons M_{(s)}$

REMEMBER THIS

The oxidation potential is reverse of the reduction potential.
Reduction potential = – Oxidation Potential

- **The electrode potential depends upon:**

i. the nature of metal and its ions,
ii. concentration of the ions in the solution, and
iii. temperature.

Cell potential or e.m.f.:

i. Galvanic cells are made up of two half-cells having different reduction potentials:
  - the electrode having higher reduction potential will have a higher tendency to gain electrons, and
  - the electrode having lower reduction potential will have higher tendency to lose the electrons.

ii. On account of this potential difference, the electrons flow from the electrode having lower reduction potential to the electrode having higher reduction potential i.e., from anode to cathode.

iii. Electric current flows in the direction opposite to that of the electron flow.

Anode (lower electrode potential)	Flow of electrons $\rightleftarrows$ Flow of current	**Cathode** (higher electrode potential)

iv. This difference in potential of two cells acts as a driving force for the cell reaction. This driving force is known as **electromotive force (e.m.f.)** or **cell potential** of a cell.

v. The cell potential of the cell is the algebraic sum of the electrode potentials for oxidation at anode and reduction at cathode. In other words, cell potential is the difference between the electrode potentials (reduction potentials) of the cathode and anode.

E_{cell} (e.m.f.) = Oxidation potential of anode + Reduction potential of cathode

$= E_{oxi}(\text{anode}) + E_{red}(\text{cathode})$

OR

E_{cell} (e.m.f.) = Reduction potential of cathode – Reduction potential of anode

$= E_{(cathode)} - E_{(anode)} = E_{(Right)} - E_{(Left)}$

vi. Cell potential is called the cell electromotive force (e.m.f.) of the cell when no current is drawn through the cell.

vii. The e.m.f. of the cell is measured with the help of a potentiometer and its unit is volt (V).

viii. The e.m.f. of cell depends upon:
  a. nature of electrodes
  b. temperature, and
  c. concentration of the solutions in two half cells.

➢ **Difference between e.m.f. and potential difference:**

No.	e.m.f.	Potential difference
i.	It is the potential difference between two electrodes in an open circuit, i.e., when there is no current flow.	It is the potential difference between two electrodes in a closed circuit, i.e., when the current is flowing through the cell.
ii.	It is the maximum voltage delivered by a cell.	It is less than e.m.f. of the cell.
iii.	It is responsible for continuous supply of current in the cell.	It is not responsible for continuous supply of current in the cell.

➢ **Standard electrode potential:**

i. As the electrode potential and cell potential depend upon the concentration of ions present in the solution and the temperature of the half-cell, therefore, to compare the electrode potential of various electrodes, fixed temperature and concentration are used.

ii. **Standard electrode potential** is the potential associated with the electrode reaction at an electrode when all solutes are 1 M and all gases are at 1 atm and at 25 °C (298 K).

iii. According to IUPAC convention, standard reduction potential is considered as standard electrode potential (E°).

MEASUREMENT OF ELECTRODE POTENTIAL

➢ **Reference electrode:**

i. The potential of individual half-cell cannot be measured. However, the e.m.f. of the cell (relative electrode potential) can be measured from the difference in the electrode potentials of two half-cells.

ii. This difficulty can be solved by connecting the electrode with a reference electrode whose potential is arbitrarily taken as zero.

iii. Standard hydrogen electrode (SHE) is used as reference electrode whose potential is assigned as zero volt. This is known as **primary reference electrode**.

iv. SHE is not a very convenient reference electrode. Therefore, several other electrodes such as calomel electrode, silver-silver chloride electrode and glass electrode are used as **secondary reference electrodes**.

v. The potential of these secondary reference electrodes can be determined accurately by using SHE as reference electrode.

vi. Therefore, the **reference electrode** is defined as an electrode whose potential is arbitrarily taken as zero or is exactly known.

➢ **Standard Hydrogen Electrode (SHE):**

An electrode in which pure and dry hydrogen gas is bubbled at 1 atm pressure around a platinised platinum plate immersed in 1 M H^+ ion solution is called ***standard hydrogen electrode (SHE)***.

i. SHE is a primary reference electrode.

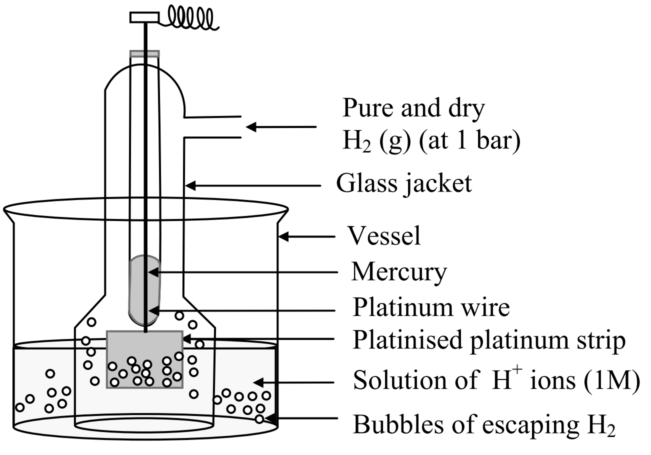
ii. Construction:

a. SHE consists of a small platinum strip coated with platinum black (to adsorb hydrogen gas) immersed in 1 M solution of H^+ ions maintained at 298K.

b. A platinum wire is welded to the platinum strip and sealed in a glass tube and connected to the outer circuit.

c. The platinum strip and glass tube are surrounded by an outer glass jacket. Hydrogen gas is bubbled into the solution through the side tube of this glass jacket at one bar pressure.

d. In this electrode, platinum is used as **inert electrode** as it does not participate in the reaction. It provides its surface for oxidation or reduction reactions and for conduction of electrons.



Standard Hydrogen Electrode (SHE)

e. Representation:

$Pt(s) \mid H_2(g) \mid H^+(aq)$, [when SHE act as anode]

$H^+(aq) \mid H_2(g) \mid Pt(s)$, [when SHE act as cathode]

where, $[H^+] = 1$ M and $P_{H_2} = 1$ bar

iii. Electrode reactions:

The following reactions occur in this half-cell depending upon whether it acts as anode or cathode.

SHE half reaction:	**Electrode Potential**
H_2 (g, 1 bar) $\rightarrow 2H^+$ (1M) + $2e^-$ (oxidation)	0.0 V (Anode)
$2H^+$ (1 M) + $2e^- \rightarrow H_2$ (g, 1 bar) (reduction)	0.0 V (Cathode)

By convention, SHE is assigned a zero potential at all temperatures.

iv. Application of SHE:

The e.m.f. of a half-cell (i.e., electrode potential) can be calculated by using SHE as reference electrode. As the electrode potential of SHE is zero, the e.m.f. of the cell is, $E_{cell} = E_{cathode} - E_{anode}$.

From this equation, unknown electrode potential of some other electrode can be calculated by substituting the value of E_{cell} and reference electrode.

E.g. a. The measured e.m.f. of the cell:

$Pt (s) \mid H_2 (g, 1 bar) \mid H^+ (aq, 1 M) \parallel Cu^{2+} (aq, 1 M) \mid Cu$

is 0.34 V which is also the value for the standard electrode potential of the half-cell corresponding to the reaction: $Cu^{2+}_{(aq)}$ (1 M) + 2 $e^- \longrightarrow Cu_{(s)}$

Positive value of the standard electrode potential indicates that Cu^{2+} ions get reduced more easily than H^+ ions. The reverse process cannot occur, i.e. H^+ ions cannot oxidise Cu (under standard conditions). Thus, Cu does not dissolve in HCl.

Note: In nitric acid, Cu is oxidized by NO_3^- ions and not by H^+ ions.

b. Similarly, the measured e.m.f. of the cell: $Zn \mid Zn^{2+}(aq, 1 M) \parallel H^+ (aq, 1 M) \mid H_2 (g, 1 bar) \mid Pt(s)$ is – 0.76 V. It corresponds to the standard electrode potential of the half-cell reaction:

$Zn^{2+}_{(aq)}$ (1 M) + $2e^- \longrightarrow Zn_{(s)}$

Negative value of the standard electrode potential indicates that H^+ ions can oxidise zinc (or zinc can reduce H^+ ions).

- **Limitations of using SHE:**

i. Pure and dry H_2 gas cannot be obtained.

ii. It is difficult to maintain the pressure of H_2 gas at exactly 1 bar.

iii. It is also difficult to maintain the concentration of H^+ ion solution at 1 M. The bubbling of H_2 gas into the solution results in the evaporation of water.

➢ **Calomel electrode:**

i. Calomel electrode is a (metal-sparingly soluble salt electrode) secondary reference electrode, which can be used conveniently.

ii. The e.m.f. of this electrode is measured by combining it with SHE, a primary reference. Hence, the name **secondary reference electrode**.

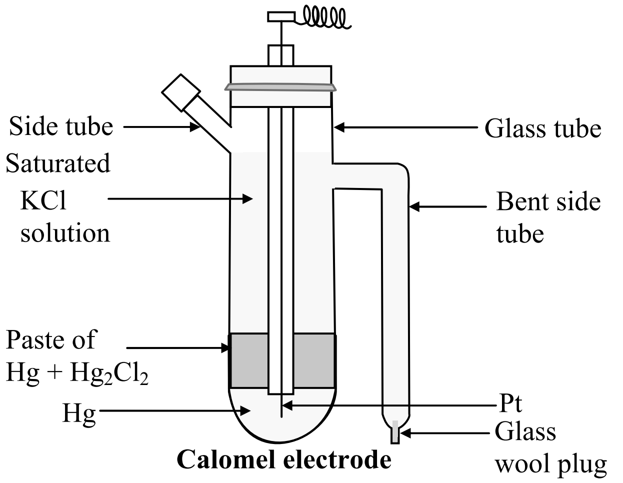
iii. Construction:

a. Calomel electrode consists of a glass tube provided with a side tube and a bent side tube.

b. This glass tube is filled with little mercury; and then a paste of mercury–mercurous chloride (calomel) and a saturated solution of potassium chloride at the top.

c. A platinum wire sealed in a glass tube is inserted in the mercury that provides electrical contact.

d. This electrode is connected to the other electrode through the bent tube to make a complete cell. This bent tube act as a salt bridge.



Calomel electrode

e. The reduction potential of calomel electrode depends upon the concentration of potassium chloride solution. The potential values based on hydrogen scale at 298 K are given in the table.

Concentration of KCl	Potential
0.1 M KCl	0.337 V
1 M KCl	0.280 V
Saturated KCl	0.242 V

f. Representation: KCl(sat) | Hg_2Cl_2(s) | Hg(*l*)

g. The electrode reaction can be any one of the following which depends on the relative electrode potential. i.e., whether the calomel electrode acts as cathode or anode in the cell.

iv. **Electrode reactions:**

If calomel electrode acts as cathode:

$Hg_2Cl_{2(s)} + 2e^- \longrightarrow 2Hg_{(l)} + 2Cl^-$ (sat.) (reduction)

If calomel electrode acts as anode:

$2Hg_{(l)} + 2Cl^-$ (sat.) $\longrightarrow Hg_2Cl_{2(s)} + 2e^-$ (oxidation)

- **Advantages of calomel electrode:**

i. It is convenient to handle, easy to construct and transport.

ii. No separate salt bridge is required for its combination with other electrode.

iii. The potential of the electrode is reproducible and remains constant.

ELECTROCHEMICAL SERIES (ELECTROMOTIVE SERIES)

➢ **Electrochemical series:**

After determining the electrode potentials of various electrodes (metal and non-metals in contact with their ions) with respect to standard hydrogen electrode, these can be arranged in decreasing order of their standard reduction potential in a series known as **electrochemical series** or **electromotive series** or **activity series** of the elements.

***Electrochemical series** is defined as the arrangement of electrodes (metal or non-metal in contact with their ions) with their half reactions in decreasing order of their standard reduction potentials.*

➢ **Convention used in the construction of electrochemical series:**

i. In electrochemical series, the electrodes are arranged in the decreasing order of their standard reduction potential and therefore, half reactions are written as reduction reactions.

ii. The electrodes with positive E° values are located above hydrogen and those with negative E° values below hydrogen. The E° value of standard hydrogen electrode is 0.00 V. Above this, positive E° value increases and below it negative E° value increases.

iii. The species on the left of the reactions are either cations of metals or hydrogen or the non-metal molecules and those on the right of the reactions are free metals or anions of non-metals.

iv. The electrodes with relatively positive value of electrode potential suggest that forward reaction (reduction reaction involving the addition of electrons) is possible. However, the electrodes with relatively negative value of electrode potential suggest that reverse reaction (oxidation reaction involving loss of electrons) is feasible.

v. When the electrode reaction occurs in reverse direction, the sign of E° changes but magnitude remains the same.

vi. The numerical value of E° remains same if the half reaction is multiplied by a numerical factor, as E° is an **intensive property.**

Electrochemical Series

Standard Reduction Electrode Potentials at 298 K

Strongest oxidizing agent

Weakest reducing agent

INCREASING STRENGTH OF OXIDIZING AGENT (Increasing tendency to accept electrons)

INCREASING STRENGTH OF REDUCING AGENT (Increasing tendency to lose electrons)

Electrode	Electrode reaction (Oxidized form + ne$^-$ → Reduced form)	Standard Electrode Reduction potential E°(V)
$F_2 \| F^-$	$F_2(g) + 2e^- \rightarrow 2F^-$ (aq)	2.87
$Co^{3+} \| Co^{2+}$	$Co^{3+} + e^- \rightarrow Co^{2+}$	1.81
$H_2O_2 \| H_2O$	$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78
$MnO_4^- \| Mn^{2+}$	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} \| Au$	$Au^{3+} + 3e^- \rightarrow Au$ (s)	1.50
$Cl_2 \| Cl^-$	Cl_2 (g) $+ 2e^- \rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-}, H^+ \| Cr^{3+}$	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2, H^+ \| H_2O$	$O_2(g) + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23
$MnO_2, H^+ \| Mn^{2+}$	$MnO_2(s) + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 \| Br^-$	Br_2 (l) $+ 2e^- \rightarrow 2Br^-$	1.09
$NO_3^-, H^+ \| NO$	$NO_3^- + 4H^+ + 3e^- \rightarrow NO$ (g) $+ 2H_2O$	0.97
$Hg^{2+} \| Hg_2^{2+}$	$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.92
$ClO^- \| Cl^-$	$ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$	0.89
$Hg^{2+} \| Hg$	$Hg^{2+} + 2e^- \rightarrow Hg$	0.85
$Ag^+ \| Ag$	$Ag^+ + e^- \rightarrow Ag$	0.80
$Hg_2^{2+} \| Hg$	$Hg_2^{2+} + e^- \rightarrow 2Hg$	0.79
$Fe^{3+} \| Fe^{2+}$	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77
O_2 (g), $H^+ \| H_2O_2$	O_2 (g) $+ 2H^+ + 2e^- \rightarrow H_2O_2$	0.68
$MnO_4^- \| MnO_4^{2-}$	$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	0.56
$I_2 \| I^-$	$I_2 + 2e^- \rightarrow 2I^-$	0.54
$Cu^+ \| Cu$	$Cu^+ + e^- \rightarrow Cu$	0.52
$Cu^{2+} \| Cu$	$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$AgCl \| Ag$	$AgCl + e^- \rightarrow Ag + Cl^-$	0.22
$Cu^{2+} \| Cu^+$	$Cu^{2+} + e^- \rightarrow Cu^+$	0.15
$AgBr \| Ag$	$AgBr + e^- \rightarrow Ag + Br^-$	0.10
$H^+ \| H_2$	**$2H^+ + 2e^- \rightarrow H_2$**	**0.00**
$Fe^{3+} \| Fe$	$Fe^{3+} + 3e^- \rightarrow Fe$	– 0.04
$Pb^{2+} \| Pb$	$Pb^{2+} + 2e^- \rightarrow Pb$	– 0.13
$Sn^{2+} \| Sn$	$Sn^{2+} + 2e^- \rightarrow Sn$	– 0.14
$Ni^{2+} \| Ni$	$Ni^{2+} + 2e^- \rightarrow Ni$	– 0.25
$Co^{2+} \| Co$	$Co^{2+} + 2e^- \rightarrow Co$	– 0.28
$Cd^{2+} \| Cd$	$Cd^{2+} + 2e^- \rightarrow Cd$	– 0.40
$Fe^{2+} \| Fe$	$Fe^{2+} + 2e^- \rightarrow Fe$	– 0.44
$Cr^{3+} \| Cr$	$Cr^{3+} + 3e^- \rightarrow Cr$	– 0.74
$Zn^{2+} \| Zn$	$Zn^{2+} + 2e^- \rightarrow Zn$	– 0.76
$H_2O \| H_{2(g)}$	$2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-$ (aq)	– 0.83
$Mn^{2+} \| Mn$	$Mn^{2+} + 2e^- \rightarrow Mn$	– 1.18
$Al^{3+} \| Al$	$Al^{3+} + 3e^- \rightarrow Al$	– 1.66
$Mg^{2+} \| Mg$	$Mg^{2+} + 2e^- \rightarrow Mg$	– 2.36
$Na^+ \| Na$	$Na^+ + e^- \rightarrow Na$	– 2.71
$Ca^{2+} \| Ca$	$Ca^{2+} + 2e^- \rightarrow Ca$	– 2.87
$Ba^{2+} \| Ba$	$Ba^{2+} + 2e^- \rightarrow Ba$	– 2.91
$K^+ \| K$	$K^+ + e^- \rightarrow K$	– 2.93
$Li^+ \| Li$	$Li^+ + e^- \rightarrow Li$	– 3.05

Weakest oxidizing agent

Strongest reducing agent

- **Applications of electrochemical series:**

A lot of useful information can be drawn from the electrochemical series. Some of them are:

i. Predicting the oxidizing and reducing ability:

a. The species on the left side of half reactions are oxidizing agents.

E° value is a measure of the tendency of the species to accept electrons and get reduced. In other words, E° value measures the strength of the substances as oxidising agents. Larger the E° value, greater is the oxidising strength. The species in the top left side of half reactions are strong oxidising agents. As we move down the table, E° value and strength of oxidising agents decreases from top to bottom.

b. The species on the right side of half reactions are reducing agents.

The half reactions at the bottom of the table with large negative E° values have a little or no tendency to occur in the forward direction as written. They tend to favour the reverse direction. It follows, that the species appearing at the bottom right side of half reactions associated with large negative E° values are the effective electron donors. They serve as strong reducing agents. The strength of reducing agents increases from top to bottom as E° values decrease.

REMEMBER THIS

Strength of oxidizing agent $\propto E^\circ$ and Strength of reducing agent $\propto \frac{1}{E^\circ}$

ii. Predicting cell e.m.f.: The standard e.m.f. of a cell can be calculated knowing the standard electrode potential values of the cathode and the anode.

$$E^o_{cell} = E^o_{cathode} - E^o_{anode}$$

E.g. The cell potential of Daniel cell can be calculated from the standard electrode potentials of copper (cathode) and zinc (anode) as follows:

$$E^o_{cell} = E^o_{cathode} - E^o_{anode} = E^o_{Cu^{2+}/Cu} - E^o_{Zn^{2+}/Zn} = 0.34V - (-0.76V) = 1.10\ V$$

iii. Predicting feasibility of reaction:

a. After calculating the standard potential of cell, the feasibility of a particular cell reaction can be predicted.

b. If E^o_{cell} = positive, the reaction is feasible and if E^o_{cell} = negative, the reaction is not feasible.

EXAMPLE – 2.1

Is it possible to store $Al_2(SO_4)_3$ in a copper container? $E^o_{Cu^{2+}/Cu} = 0.34\ V$, $E^o_{Al^{3+}/Al} = -1.66\ V$

Solution:

The required reaction is : $2Al^{3+}_{(aq)} + 3Cu_{(s)} \longrightarrow 2Al_{(s)} + 3Cu^{2+}_{(aq)}$

The feasibility of the above reaction can be predicted by determining the E^o_{cell} value.

Electrode reactions:

At Cathode: $2Al^{3+}_{(aq)} + 6e^- \longrightarrow 2Al_{(s)}$ $E^\circ = -1.66\ V$

At Anode: $3Cu_{(s)} \longrightarrow 3Cu^{2+}_{(aq)} + 6e^-$ $E^\circ = 0.34\ V$

$$E^o_{cell} = E^o_{cathode} - E^o_{anode} = -1.66 - 0.34 = \mathbf{-2.00\ V}$$

The negative value of cell e.m.f. suggests that the reaction is not feasible.

Hence, it is possible to store $Al_2(SO_4)_3$ in a copper container.

iv. Predicting whether a metal will displace another metal from its salt solution or not:

a. The metals near the bottom of electrochemical series are strong reducing agents. These metals reduce the metal ions to metal and themselves get oxidized to metal ions.

b. Therefore, a metal present in the series can displace another metal present above in the electrochemical series from its salt solution.

c. Consider the following example:
If iron nails are placed in copper sulphate solution, iron displaces copper (present above the iron in electrochemical series) from copper sulphate solution. As a result, copper is deposited on nail and iron gets oxidized to ferrous ions. That is why the solution becomes green in colour due to the formation of ferrous sulphate. This change can be represented by the following reaction:

$$Fe_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$$

REMEMBER THIS

A reducing agent can reduce the oxidizing agent located above it in the electrochemical series.
An oxidizing agent can oxidise the reducing agent located below it in the electrochemical series.

EXAMPLE – 2.2

Half reactions along with standard reduction potentials are given below:

Half reaction	E° value (V)
$Au^{3+} + 3e^- \to Au$	+1.50
$Pb^{2+} + 2e^- \to Pb$	–0.13
$Co^{2+} + 2e^- \to Co$	–0.28
$Zn^{2+} + 2e^- \to Zn$	–0.76
$Ca^{2+} + 2e^- \to Ca$	–2.87
$Ba^{2+} + 2e^- \to Ba$	–2.91

Based on the E° values, predict which of the following ions in aqueous solution can oxidize only one metal from the given list.

(A) Au^{3+} (B) Zn^{2+} (C) Pb^{2+} (D) Ca^{2+}

Solution:

Greater E° value indicates greater tendency of the species on the left to accept electrons and undergo reduction (that is, they act as oxidizing agents).

In general, any oxidizing agent (the species on the left of half reaction) can oxidize any reducing agent (the species on the right side of half reaction) that appears below it but cannot oxidize the species located above it in the electrochemical series.

Hence, from the given list, Ca^{2+} can oxidize only Ba.

Ans: (D)

v. Predicting whether a non-metal will displace another non-metal from its salt solution:

A non-metal having higher value of reduction potential will displace another non-metal with lower reduction potential (i.e. present below in the series).

E.g. Cl_2 can displace bromine and iodine from bromides and iodides, respectively.

$$Cl_2 + 2KI \longrightarrow 2KCl + I_2$$

vi. Product of electrolysis:

a. During electrolysis, if two or more types of positive and negative ions are present in solution, certain ions are discharged at electrode in preference to others.

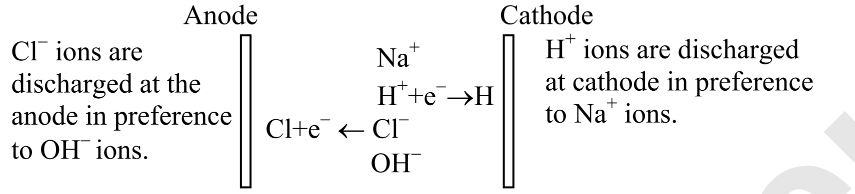
b. In such competition, the ion which is stronger oxidizing agent (having high value of standard reduction potential) is discharged first at the cathode.

$$\left(\begin{array}{c}\text{Stronger}\\\text{oxidizing}\\\text{agent}\end{array}\right) \xrightarrow[\text{Decreasing order of deposition at cathode}]{Au^{3+}, Ag^+, Cu^{2+}, H^+, Fe^{2+}, Zn^{2+}, Al^{3+}, Mg^{2+}, Na^+, Ca^{2+}, K^+} \left(\begin{array}{c}\text{Weaker}\\\text{oxidizing}\\\text{agent}\end{array}\right)$$

c. Similarly, the anion which is stronger reducing agent (having low value of standard reduction potential) is discharged first at the anode.

$$\begin{pmatrix}\text{Stronger}\\ \text{reducing}\\ \text{agent}\end{pmatrix} \xrightarrow[\text{Decreasing order of deposition at anode}]{I^-,\ Br^-,\ Cl^-,\ OH^-,\ NO_3^-,} \begin{pmatrix}\text{Weaker}\\ \text{reducing}\\ \text{agent}\end{pmatrix}$$

d. **E.g.** An aqueous solution of NaCl containing Na^+, Cl^-, H^+ and OH^- ions on electrolysis liberates hydrogen gas at cathode and chlorine gas at anode.



The product of electrolysis of aqueous solution of NaCl

For your knowledge

- **Latimer diagram:**

i. Latimer or reduction potential diagrams show the standard reduction potentials connecting various oxidation states of an element.

ii. Latimer diagram showing relative stabilities of different oxidation states are given below:

a. Acid Medium: $Fe \xrightarrow{E^\circ_{Fe/Fe^{2+}} = +0.44\,V} Fe^{2+} \xrightarrow{E^\circ_{Fe^{2+}/Fe^{3+}} = -0.79\,V} Fe^{3+}$

b. Alkaline Medium: $Fe \xrightarrow{E^\circ_{Fe/Fe^{2+}} = +0.88\,V} Fe^{2+} \xrightarrow{E^\circ_{Fe^{2+}/Fe^{3+}} = +0.56\,V} Fe^{3+}$

c. In acidic medium, Fe^{2+} state is more stable as represented by the respective standard electrode potentials. The positive value of $E^\circ_{Fe/Fe^{2+}}$ indicates that forward reaction is feasible, therefore, iron will dissolve in acid medium to form Fe^{2+} ion, but $E^\circ_{Fe^{2+}/Fe^{3+}}$ is negative; hence it exists in Fe^{2+} state which is more stable than Fe^{3+} state in acidic medium.

iii. However, in Latimer diagram, if the potential on right of the species is more positive or less negative than that on the left, then the species undergoes disproportionation (i.e., oxidation-reduction simultaneously).

E.g. Cl_2 undergoes disproportionation into Cl^- and ClO^- ions in alkaline medium.

Alkaline Medium: $Cl^- \xrightarrow{E^\circ_{Cl^-/Cl_2} = -1.36V} \frac{1}{2}Cl_2 \xrightarrow{E^\circ_{Cl_2/ClO^-} = +0.40V} ClO^-$

Therefore, $Cl_2 + 2OH^- \longrightarrow Cl^- + ClO^- + H_2O$

RELATION BETWEEN GIBBS ENERGY CHANGE AND E.M.F. OF A CELL

- **Gibbs energy change and e.m.f. of a cell:**

i. In a galvanic cell, electrical energy is produced due to the flow of current and this electrical energy can be used to perform work.

ii. The electrical work done in a galvanic cell is the product of electricity (charge) passed and the cell potential. Maximum work can be obtained if the charge is passed reversibly.

∴ Electrical work = amount of charge × cell potential

∴ $W_{max} = nF \times E_{cell}$

$W_{max} = nFE_{cell}$ ….(1)

where, n = number of moles of electrons transferred and is equal to the valence of ions participating in the cell reaction, F = Faraday constant ($96485\ C\ mol^{-1} \approx 96500\ C\ mol^{-1}$) and nF (= Q coulombs) gives the amount of charge passed.

iii. According to thermodynamics, the maximum work that can be derived from a chemical reaction is equal to the decrease in Gibbs energy of the reaction.

i.e., $W_{max} = -\Delta_rG$ ….(2)

iv. From equation (1) and (2),

$-\Delta_rG = nFE_{cell}$ **OR** $\Delta_rG = -nFE_{cell}$ ….(3)

v. If the concentration of all the reacting species is unity, then

$E_{cell} = E^o_{cell}$ ….(4)

$\therefore$ $\Delta_rG° = -nFE^o_{cell}$ ….(5)

vi. From the value of E^o_{cell}, $\Delta_rG°$ (standard Gibbs energy) of the reaction can be calculated which is an important thermodynamic quantity.

EXAMPLE – 2.3

Calculate the standard Gibbs energy change for the following reaction at 298 K:

$2Au_{(s)} + 3Ca^{2+}(1.0\text{ M}) \longrightarrow 2Au^{3+}(1.0\text{ M}) + 3Ca_{(s)}$

Given: $E°(Ca^{2+} | Ca) = -2.87\text{ V}$ and $E°(Au^{3+} | Au) = 1.50\text{ V}$

Solution:

Overall cell reaction: $2Au_{(s)} + 3Ca^{2+} \longrightarrow 2Au^{3+} + 3Ca_{(s)}$

The half-cell reactions are

Anode (oxidation): $2Au_{(s)} \longrightarrow 2Au^{3+} + 6e^-$

Cathode (reduction): $3Ca^{2+} + 6e^- \longrightarrow 3Ca_{(s)}$

Therefore, Au is the anode, Ca is the cathode and n = 6.

$$E^o_{cell} = E^o_{cathode} - E^o_{anode}$$

$$= E^o_{Ca^{2+}/Ca} - E^o_{Au^{3+}/Au}$$

$$= -2.87\text{ V} - 1.50\text{ V} = -4.37\text{ V}$$

Now, $\Delta_rG° = -nF\,E^o_{cell}$

$\therefore$ $\Delta_rG° = -(6)(96{,}500\text{ C/mol})(-4.37\text{ J/C})$ [since, 1 V = 1 J/C]

$= 2.53 \times 10^6$ J/mol

$= \mathbf{2.53 \times 10^3}$ **kJ/mol**

- **Standard cell potential and equilibrium constant:**

i. The standard Gibbs energy change of a reaction is related to equilibrium constant (K) as:

$\Delta_rG° = -RT\ln K$ ….(6)

ii. From equation (5) and (6);

$-nFE^o_{cell} = -RT\ln K$

$$E^o_{cell} = \frac{RT}{nF}\ln K = \frac{2.303\,RT}{nF}\log_{10}K = \frac{0.059}{n}\log_{10}K \text{ at 298 K}$$

where, $R = 8.314\text{ J K}^{-1}\text{ mol}^{-1}$, T = 298 K and $F = 96500\text{ C mol}^{-1}$

Note: The value of K gives the extent of the cell reaction. If the value of K is large, the reaction proceeds to larger extent.

Smart tip - 1

Equilibrium constant, $K = 10^{(n \times E°/0.059)}$ at 298 K.

When **n = 1** , $K = 10^{(E°/0.059)}$ OR $K = 10^{(\mathbf{16.9} \times E°)}$

When **n = 2** , $K = 10^{(2E°/0.059)}$ OR $K = 10^{(\mathbf{33.8} \times E°)}$

When **n = 3** , $K = 10^{(3E°/0.059)}$ OR $K = 10^{(\mathbf{50.8} \times E°)}$

EXAMPLE – 2.4

The value of equilibrium constant for the following reaction at 298 K is _______.

$Cu_{(s)} + 2Ag^{+}_{(aq)} \rightleftharpoons 2Ag_{(s)} + Cu^{2+}_{(aq)}$

Given that $E^{\circ}_{\left(Ag^{+}/Ag\right)} = 0.80$ V and $E^{\circ}_{\left(Cu^{2+}/Cu\right)} = 0.34$ V

(A) $10^{0.46}$ (B) $10^{15.5}$

(C) $10^{7.7}$ (D) $10^{3.7}$

Solution:

Reaction: $Cu_{(s)} + 2Ag^{+}_{(aq)} \rightleftharpoons 2Ag_{(s)} + Cu^{2+}_{(aq)}$

Ag electrode is cathode, Cu electrode is anode and n = 2.

$$E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode}$$

$$= E^{o}_{Ag^{+}/Ag} - E^{o}_{Cu^{2+}/Cu}$$

$$= (0.80 - 0.34) \text{ V}$$

$$E^{o}_{cell} = 0.46 \text{ V}$$

Now, $E^{o}_{cell} = \dfrac{0.059}{n} \log_{10} K$ at 298 K

OR

$K = 10^{(33.8 \times E^{\circ})}$ (Refer ***Smart tip - 1***)

$\therefore$ $K = 10^{(33.8 \times 0.46)}$

$K = 10^{15.5}$

Ans: (B)

NERNST EQUATION AND ITS APPLICATIONS

➢ **Derivation of Nernst equation:**

i. The electrode potential of a cell depends upon the nature of electrode, temperature and concentrations of the ions in the solutions.

ii. **Walther H. Nernst** (1864 – 1941) gave the relationship between electrode potential and the concentration of electrolyte solutions (which may be different from unity), known as **Nernst equation**.

iii. For a general reaction of the type, $aA + bB \rightleftharpoons cC + dD$, the free energy change is given by the equation,

$$\Delta G = \Delta_r G^{o} + RT \ln \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad \text{....(1)}$$

where $K = \dfrac{[C]^c [D]^d}{[A]^a [B]^b}$

and, $\Delta G = -\, nFE_{cell}$ ....(2)

$\Delta_r G^{o} = -\, nFE^{o}_{cell}$ ....(3)

Substituting (2) and (3) in equation (1),

$$\therefore \quad -\,nFE_{cell} = -\,nFE^{o}_{cell} + RT \ln \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

$$\therefore \quad E_{cell} = E^{o}_{cell} - \frac{RT}{nF} \ln \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad \text{....(4)}$$

By converting the natural logarithm to log to the base 10,

$$\mathbf{E_{cell} = E^{o}_{cell} - \frac{2.303\, RT}{nF} \log_{10} \frac{[C]^c [D]^d}{[A]^a [B]^b}} \quad \text{....(5)}$$

iv. Substituting the values, R = 8.314 J K^{-1} mol^{-1}, T = 298 K and F = 96500 C mol^{-1}

$$E_{cell} = E^{o}_{cell} - \frac{0.059}{n} \log_{10} \frac{[C]^c[D]^d}{[A]^a[B]^b} \qquad ....(6)$$

$$\therefore \quad \mathbf{E_{cell} = E^{o}_{cell} - \frac{0.059}{n} \log_{10} \frac{[Products]}{[Reactants]}} \qquad ....(7)$$

Equation (4), (5), (6) and (7) are the different forms of Nernst equation.

REMEMBER THIS

If the concentration of reactants and products are 1 M each, then

$$E_{cell} = E^{o}_{cell} - \frac{0.059}{n} \log_{10} \frac{1}{1} = E^{o}_{cell}$$

∴ The E_{cell} is equal to standard cell potential (E^{o}_{cell}) at unit concentration of reactants and products.

Smart tip - 2

According to Nernst equation, $E_{cell} = E^{o}_{cell} - \frac{RT}{nF} \ln \frac{[Products]}{[Reactants]}$

i.	Ratio $\frac{[Products]}{[Reactants]} < 1$	[Products] < [Reactants]	$\ln \frac{[Products]}{[Reactants]}$ is **negative**	$E_{cell} > E^{o}_{cell}$
ii.	Ratio $\frac{[Products]}{[Reactants]} > 1$	[Products] > [Reactants]	$\ln \frac{[Products]}{[Reactants]}$ is **positive**	$E_{cell} < E^{o}_{cell}$

EXAMPLE – 2.5

Consider the following cell reaction:

Sn^{4+} (1.5 M) + $Zn_{(s)} \longrightarrow Sn^{2+}$ (0.5 M) + Zn^{2+} (2 M)

i. Find out the relation between $E_{cell} = E^{o}_{cell}$ for this cell.

ii. Will the value of E_{cell} increase or decrease if the concentration of Sn^{2+} is increased?

Solution:

i. For the given cell reaction:
[Products] = 0.5 × 2 = 1 M and [Reactants] = 1.5 M
So, [Products] < [Reactants]
Hence, $\mathbf{E_{cell} > E^{o}_{cell}}$

ii. For the given cell reaction:
Sn^{2+} is a product. If the concentration of Sn^{2+} is increased, then [Products] > [Reactants].
Hence, the value of E_{cell} will decrease if the concentration of Sn^{2+} is increased.

➢ **Applications of Nernst equation:**

Nernst equation has the following applications:

i. Calculation of electrode potential:

a. For a reduction reaction; $M^{n+}_{(aq)} + ne^- \longrightarrow M_{(s)}$

The electrode potential by Nernst equation can be written as: $E_{(M^{n+}/M)} = E^{\circ}_{(M^{n+}/M)} - \frac{RT}{nF} \ln \frac{[M]}{[M^{n+}]}$

But concentration of solid is taken as unity, i.e., [M] = 1

$$\therefore \quad E_{(M^{n+}/M)} = E^{\circ}_{(M^{n+}/M)} - \frac{RT}{nF} \ln \frac{1}{[M^{n+}]}$$

By converting the natural logarithm to log to the base 10 and substituting the values of R, F and at T = 298 K,

the equation becomes: $\mathbf{E_{(M^{n+}/M)} = E^{\circ}_{(M^{n+}/M)} - \frac{0.059}{n} \log_{10} \frac{1}{[M^{n+}]}}$

From this equation, electrode potential can be calculated.

b. For example, the electrode potential for the electrodes of Daniel cell at a given concentration of Cu^{2+} and Zn^{2+} ions can be calculated by the Nernst equation.

c. **For Cathode:** $E_{(Cu^{2+}/Cu)} = E°_{(Cu^{2+}/Cu)} - \frac{0.059}{n} \log_{10} \frac{1}{[Cu^{2+}]}$

d. **For Anode:** $E_{(Zn^{2+}/Zn)} = E^{\circ}_{(Zn^{2+}/Zn)} - \frac{0.059}{n} \log_{10} \frac{1}{[Zn^{2+}]}$

Smart tip - 3

For hydrogen gas electrode at 298 K:

Electrode reaction: $2H^+ + 2e^- \longrightarrow H_{2(g)}$

Using Nernst equation,

$$E_{H^+/H_2} = E° - \frac{0.059}{2} \times \log_{10} \frac{P_{H_2}}{[H^+]^2} = -\frac{0.059}{2} \times \log_{10} \frac{P_{H_2}}{[H^+]^2} \quad [\text{Since, } E° = 0]$$

	Condition	Modified Nernst equation
i.	Pressure of H_2 is 1 bar (or 1 atm) and $[H^+] \neq 1$ M	$E_{H^+/H_2} = \frac{0.059}{2} \times \log_{10} [H^+]^2$ $= 0.059 \times \log_{10} [H^+]$ $= 0.059 \times (-pH)$
ii.	$[H^+]$ is 1 M and pressure of $H_2 \neq 1$ bar (or 1 atm)	$E_{H^+/H_2} = -\frac{0.059}{2} \times \log_{10} P_{H_2}$

ii. **Calculation of cell potential:**

a. Cell potential of a cell having varying concentration of the solution can be calculated with the help of Nernst equation.

b. For example, consider a Daniel cell; $Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}$

c. The cell reaction is, $Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$

d. The cell potential for the above reaction (at 298 K) is:

$$E_{cell} = E^o_{cell} - \frac{0.059}{2} \log_{10} \frac{[Zn^{2+}][Cu]}{[Zn][Cu^{2+}]}$$

The concentration of solids is taken as unity, i.e., [Zn] = [Cu] = 1.

e. Therefore, the equation reduces to:

$$E_{cell} = E^o_{cell} - \frac{0.059}{2} \log_{10} \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

From this equation, E_{cell} can be calculated where, $E^o_{cell} = E^o_{cathode} - E^o_{anode}$

Note: In Nernst equation, electrode potential is always taken as **reduction potential**.

EXAMPLE – 2.6

Write Nernst equation for the following cells at 298 K:

i. $Cu(s) \mid Cu^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$ ii. $Al(s) \mid Al^{3+}(aq) \parallel Ni^{2+}(aq) \mid Ni(s)$

Solution:

Nernst equation is : $E_{cell} = E^o_{cell} - \frac{0.059}{n} \log_{10} \frac{[Products]}{[Reactants]}$ at 298 K

i. $Cu(s) \mid Cu^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$

Cell reaction: $Cu_{(s)} + 2\,Ag^+_{(aq)} \longrightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

Here n = 2;

∴ Nernst equation for this cell is: $E_{cell} = E^o_{cell} - \frac{0.059}{2} \log_{10} \frac{[Cu^{2+}]}{[Ag^+]^2}$ at 298 K

ii. $Al(s) \mid Al^{3+}(aq) \parallel Ni^{2+}(aq) \mid Ni(s)$

Cell reaction: $2Al_{(s)} + 3Ni^{2+}_{(aq)} \longrightarrow 3Ni_{(s)} + 2Al^{3+}_{(aq)}$

Here $n = 6 = 3 \times 2e^-$ (for Ni^{2+}) $= 2 \times 3e^-$ (for Al^{3+});

$\therefore$ Nernst equation for this cell is: $E_{cell} = E^o_{cell} - \frac{0.059}{6} \log_{10} \frac{[Al^{3+}]^2}{[Ni^{2+}]^3}$ at 298 K

iii. Calculation of concentration of a solution of half-cell:

The concentration of a half-cell can be calculated if the concentration of the other half-cell, E^o_{cell} and E_{cell} at that concentration are known by using the Nernst equation

iv. Calculation of equilibrium constant from Nernst equation:

a. At equilibrium, the cell potential is zero as the electrode potentials of the two electrodes become equal.

b. For example, consider the following reaction; $Zn_{(s)} + Cu^{2+}_{(aq)} \rightleftharpoons Zn^{2+}_{(aq)} + Cu_{(s)}$

c. The cell potential by Nernst equation can be written as: $E_{cell} = E^o_{cell} - \frac{0.059}{n} \log_{10} \frac{[Zn^{2+}]}{[Cu^{2+}]}$ at 298 K

d. At equilibrium, $E_{cell} = 0$ and the equilibrium constant, $K_C = \frac{[Zn^{2+}]}{[Cu^{2+}]}$

$\therefore$ $E^o_{cell} = \frac{0.059}{n} \log_{10} K_C$ at 298 K

From the above equation, equilibrium constant can be calculated.

e. For solubility product constant (K_{sp}), the equation is

$E^o_{cell} = \frac{0.059}{n} \log_{10} K_{sp}$ at 298 K

v. To predict the feasibility of a reaction:

a. The Gibbs energy change and electrode potential of a reaction are related by the equation, $-\Delta_r G = nFE_{cell}$.

b. E_{cell} can be calculated by the Nernst equation.

c. A cell reaction is feasible or spontaneous, if $\Delta_r G$ is negative or E_{cell} is positive.

For your knowledge

➢ **Concentration cells:**

- *When two half-cells of same type are connected with only the concentration of electrolyte being different, the cell thus formed is called* ***electrolyte concentration cell.***

E.g. $Cu \mid CuSO_4\,(aq)\,(C_1) \parallel CuSO_4\,(aq)\,(C_2) \mid Cu$ **where $C_2 > C_1$**

$\therefore$ $E_{cell} = \frac{RT}{nF} \ln \frac{C_2}{C_1} = \frac{0.059}{n} \log \frac{C_2}{C_1}$ at 298 K

The electrode reactions will be:

At Cathode: $Cu^{2+}\,(C_2) + 2e^- \longrightarrow Cu_{(s)}$

At Anode: $Cu_{(s)} \longrightarrow Cu^{2+}\,(C_1) + 2e^-$

Overall reaction: $Cu^{2+}\,(C_2) \longrightarrow Cu^{2+}\,(C_1)$

Thus, there is no net chemical reaction. It only involves the transfer of electrolyte from one half cell to the other.

E_{cell} will be positive till $C_2 > C_1$. When $C_1 = C_2$, $E_{cell} = 0$ and the cell will stop operating.

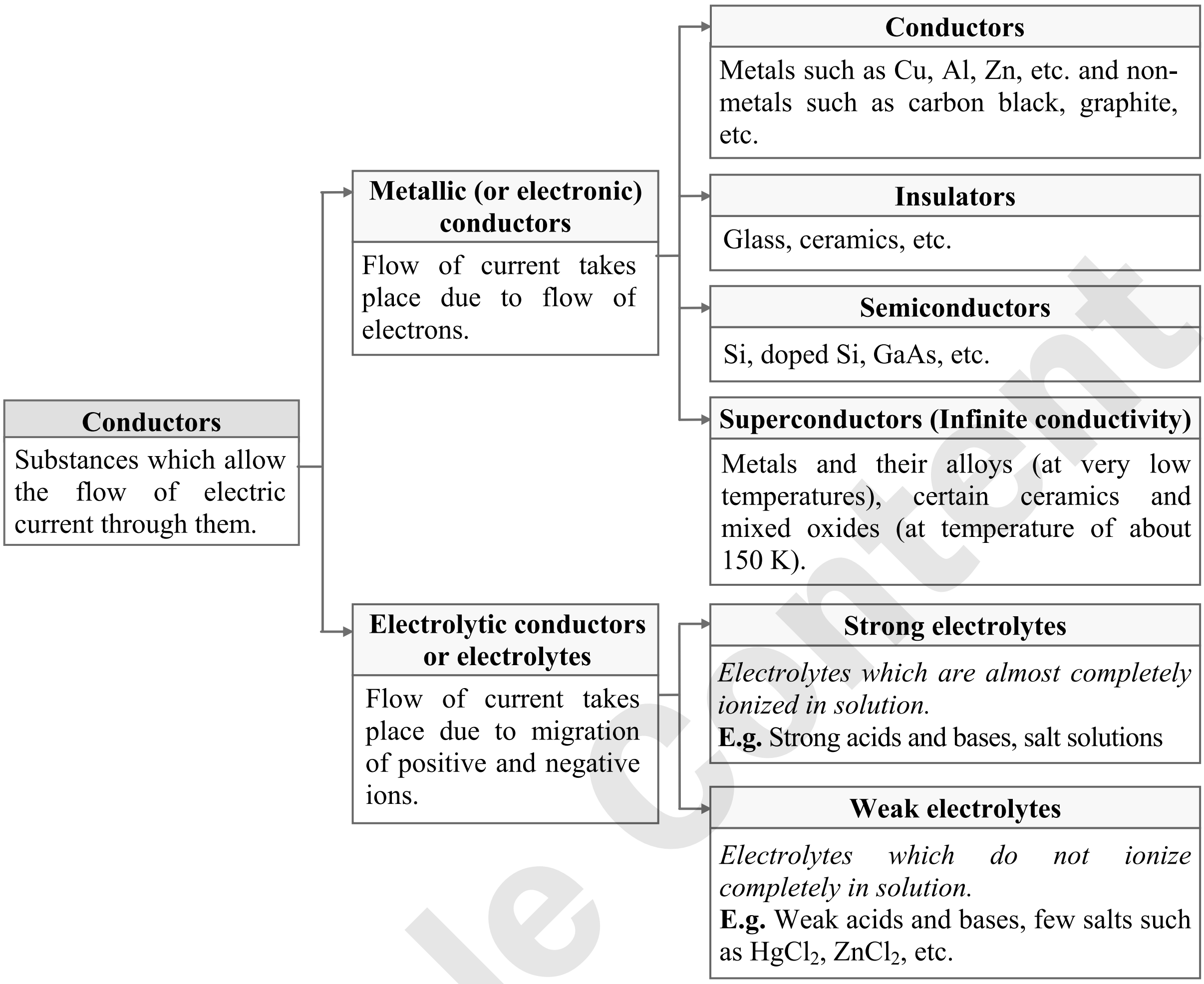
- *When two electrodes having different concentrations are dipped in the same electrolyte, the cell thus formed is called* ***electrode concentration cell.***

E.g. Pt, H_2 (P_1 bar) | H^+ (C) | H_2 (P_2 bar), Pt **where $P_1 > P_2$**

$\therefore$ $E_{cell} = \frac{0.059}{n} \log \frac{P_1}{P_2}$ at 298 K

CONDUCTANCE IN ELECTROLYTIC SOLUTIONS

- **Conductors:**



Conductors
Substances which allow the flow of electric current through them.

Metallic (or electronic) conductors
Flow of current takes place due to flow of electrons.

Conductors
Metals such as Cu, Al, Zn, etc. and non-metals such as carbon black, graphite, etc.

Insulators
Glass, ceramics, etc.

Semiconductors
Si, doped Si, GaAs, etc.

Superconductors (Infinite conductivity)
Metals and their alloys (at very low temperatures), certain ceramics and mixed oxides (at temperature of about 150 K).

Electrolytic conductors or electrolytes
Flow of current takes place due to migration of positive and negative ions.

Strong electrolytes
Electrolytes which are almost completely ionized in solution.
E.g. Strong acids and bases, salt solutions

Weak electrolytes
Electrolytes which do not ionize completely in solution.
E.g. Weak acids and bases, few salts such as $HgCl_2$, $ZnCl_2$, etc.

- **Difference between Electronic and Electrolytic conduction:**

No.	Electronic or metallic conduction	Electrolytic conduction
i.	Conduction takes place by movement of electrons.	Conduction takes place by movement of positive and negative ions.
ii.	Conduction does not involve the transfer of any matter.	Conduction involves the transfer of matter.
iii.	Chemical properties of the substance remain same after conduction.	Chemical properties of the substance change after conduction due to the decomposition of electrolyte.
iv.	Metallic conduction decreases with increase in temperature.	Electrolytic conduction increases with increase in temperature.

- **Non-electrolytes:**

__Non-electrolytes__ are the substances which do not conduct electricity either in molten state or in aqueous solutions.

These substances are covalent solids that do not dissociate into the ions in aqueous solutions and the conductivity of their solutions is equal to the conductivity of water.
E.g. Glucose, sugar, sucrose, urea, ethyl alcohol, etc. are non-electrolytes.

- **Factors affecting electronic or metallic conduction:**
The conductivity of metals depends upon the following factors:

i. the nature and structure of the metal
ii. the number of valence electrons per atom
iii. temperature (it decreases with increase of temperature).

- **Factors affecting electrical conductivity of electrolytic solutions:**
  The conductivity of electrolytic solutions depends upon the following factors:

i. Nature of electrolyte:
- a. Conductivity of electrolyte $\propto$ Number of ions present in the solution.
  Greater the number of ions, greater is the conductivity of solution.
- b. Strong electrolytes ionize almost completely in solution and their solution has greater conductivity than the solution of weak electrolytes which are weakly ionized.

ii. Nature of solvent:
- a. Ionization of electrolytes takes place to a greater extent in polar solvents.
- b. In other words, greater the polarity of the solvent, more the ionization and greater the conductivity.
- c. For strong electrolytes, degree of ionization (α) is almost equal to 1 and for weak electrolytes, it is always less than one.

iii. Size of the ions and their solvation:
If the ions of electrolytes are highly solvated, the size of the ions increases, resulting in the decrease in conductivity of the solution.

iv. Concentration of solution:
- a. Ionization of electrolyte is inversely proportional to the concentration of the solution.
- b. On dilution $\rightarrow$ Concentration of solution decreases $\rightarrow$ ionization increases,
  thus, increasing the conductivity of the solution.

v. Temperature:
- a. Temperature of solution increases $\Rightarrow$ Ionization of electrolyte in solution increases $\Rightarrow$ Conductivity of the solution increases
- b. Increase in temperature results in higher molecular velocities which overcome the forces of attraction between the ions, resulting in greater ionization.

- **Electrolytic conduction (Conductance in electrolytic solutions):**

i. When the electrodes, dipped into an electrolytic solution are connected and switch is 'on' position, the solution conducts electrical current through them by the movement of ions to the electrodes and electrons flow from negative electrode to positive electrode. i.e., the direction of the current flow is opposite to that of the electron flow.
ii. The power of electrolytes to conduct electrical currents is termed as conductance or conductivity.
iii. Electrolytic solutions also obey Ohm's law like metallic conductors.
iv. Ohm's Law:

> **Statement:** *The current flowing through a conductor is directly proportional to the potential difference across it.*

Therefore, $I \propto V$ and $I = \frac{V}{R}$ or $V = IR$

where, I – Current (in amperes),
V– Potential difference applied across the conductor (in volts),
R – Constant of proportionality, known as resistance of the conductor (in ohms, Ω).

- **Some basic terms:**
  Some commonly used basic terms in electrochemistry are listed below:

- **Resistance:**

> ***Resistance*** *gives the measure of obstruction to the flow of current.*

Resistance, R of a conductor is directly proportional to its length, (*l*) and inversely proportional to the area of cross section, (a).

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

$$\mathbf{R = \rho \frac{\mathit{l}}{a}}$$

where ρ (Greek, rho) – proportionality constant, known as specific resistance or resistivity (IUPAC recommended term). Its value depends upon the material (nature) of the conductor.

Unit of resistance: ohm (Ω), in SI base units it is equal to $\left(\text{kg}\,\text{m}^2\right)/\left(\text{S}^3\text{A}^2\right)$.

- **Specific resistance or Resistivity:**

  $\rho = R\dfrac{a}{l}$

  If $l = 1$ m and $a = 1$ m^2, then, $\rho = R$. Thus,

> ***Resistivity*** *is the resistance of the conductor of 1 m length and having an area of cross section of 1 m^2.*
>
> **OR**
>
> ***Resistivity*** *of a conductor is the resistance in ohms offered by one metre cube of it to the passage of electricity.*

Unit of resistivity: ohm m (SI unit) or ohm cm (C.G.S unit).

- **Conductance:**

> ***Conductance*** *is a measure of the ease with which current flows through a conductor.*

Therefore, Conductance, $G = \dfrac{1}{R}$ or $G = \dfrac{a}{\rho\, l}$

$\mathbf{G = \kappa\dfrac{a}{\mathit{l}}}$

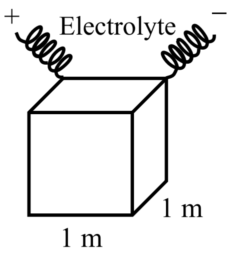
where κ (Greek, kappa) − proportionality constant, known as specific conductance or conductivity (IUPAC recommended term) and it is inverse of resistivity.
Unit of conductance: ohm^{-1} (or mho, Ω^{-1}), SI unit is siemens (S), 1 S = 1 Ω^{-1}.
Note: The unit siemens is named after Sir William Siemens, a noted electrical engineer.

- **Specific conductance or Conductivity:**

  $\kappa = G\,\dfrac{l}{a}$

  If, $l = 1$ m and $a = 1$ m^2, then, $\kappa = G$. Thus,

> ***Conductivity*** *is defined as the conductance of a solution (conductor) of 1 m length and having 1 m^2 area of cross-section.*
>
> ***OR***
>
> ***Conductivity*** *is defined as the conductance of one metre cube of a solution of an electrolyte.*



Conductance of cube = Conductivity

Diagrammatic illustration of conductivity

Unit of conductivity: ohm^{-1} m^{-1} or Ω^{-1} m^{-1} or S m^{-1} (SI unit), Ω^{-1} cm^{-1} (C.G.S unit)
The magnitude of conductivity depends upon the nature of material.

- **Molar conductance or Molar conductivity:**
  Conductivity of a solution depends upon the number of ions present in the solution, i.e., the concentration of the solution. Therefore, to compare the conductivities of different solutions, they must have same concentration.

> ***Molar conductivity*** *is defined as the conductance of all ions produced by one mole (one gram-molecular weight) of an electrolyte when dissolved in a certain volume V in m^3.*

Molar conductivity, $\wedge_m = \kappa V$

Since $V = \dfrac{1}{C}$, we have

$\wedge_m\ (\text{S m}^2\ \text{mol}^{-1}) = \dfrac{\kappa\left(\text{S m}^{-1}\right)}{C\left(\text{mol m}^{-3}\right)}$ **OR** $\wedge_m\ (\Omega^{-1}\ \text{cm}^2\ \text{mol}^{-1}) = \dfrac{\kappa\left(\Omega^{-1}\text{cm}^{-1}\right)\times\left(1000\ \text{cm}^3\ \text{L}^{-1}\right)}{\text{Molarity}\left(\text{mol L}^{-1}\right)}$

Now, $\mathbf{\wedge_m = \dfrac{G\,\mathit{l}\,V}{a}}$

Therefore, **Molar conductivity** of a solution at a given concentration is the conductance of the volume V of solution containing one mole of electrolyte kept between two electrodes with unit area of cross section (a) and distance of unit length (l).
Unit of molar conductivity: S m^2 mol^{-1} (SI unit) or S cm^2 mol^{-1} (C.G.S unit)

- **Equivalent conductivity:**

> ***Equivalent conductivity*** *is the conducting power of all the ions produced by dissolving one gram equivalent of an electrolyte in solution.*

Equivalent conductivity ($\wedge_e$) ($S\ m^2\ eq^{-1}$) = $\dfrac{\kappa\left(S\ m^{-1}\right)}{C_{eq}\left(eq\,m^{-3}\right)}$ or $\dfrac{\kappa\left(S cm^{-1}\right)\times 1000\ cm^3\ L^{-1}}{\text{Normality}\left(eq\,L^{-1}\right)}$

where, κ is specific conductance and C_{eq} is the concentration in gram equivalent per litre (or normality)

Unit of equivalent conductivity:

$ohm^{-1}\ m^2$ (gram equivalent)$^{-1}$ or $S\ m^2\ eq^{-1}$ (SI unit) or $S\ cm^2\ eq^{-1}$ (C.G.S unit)

Nowadays, equivalent conductivity is replaced by molar conductivity.

Note: Dimensions of electrochemical cells are in cm, for practical purposes. Hence, measurements are made in units involving cm and then converted to units involving m for reporting as per SI convention.

EXAMPLE – 2.7

The conductivity of 0.1 mol L^{-1} KCl solution is 1.41×10^{-3} S cm^{-1}. What is its molar conductivity (in S cm^2 mol^{-1})? **[TS EAMCET (Med.) 2015]**

(A) 14.1 (B) 1.41

(C) 1410 (D) 141

Solution:

Molar conductivity ($\wedge_m$)

$= \dfrac{1000\kappa}{C}$

$= \dfrac{1000\ cm^3 L^{-1} \times 1.41 \times 10^{-3}\ S\ cm^{-1}}{0.1\ mol\,L^{-1}}$

$= 14.1\ S\ cm^2\ mol^{-1}$

Ans: (A)

MEASUREMENT OF CONDUCTIVITY

i. The conductivity of a solution can be determined from the resistance measurements by Wheatstone bridge.

ii. However, in measuring the resistance of ionic solutions, two problems might be faced:
  a. DC current changes the composition of solution. This can be overcome by using AC current.
  b. A solution cannot be connected to the bridge like a wire. This problem is solved by using specially designed vessel called ‘conductivity cell’.

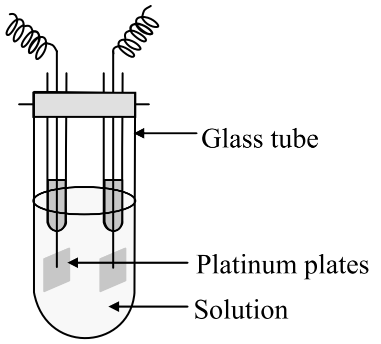
- **Conductivity cell:**

i. A conductivity cell basically consists of two platinum electrodes coated with platinum black.

ii. These platinum plates have area of cross section ‘a’ and are separated by distance ‘*l*’.

iii. The conductivity cell is filled with the solution whose resistance (R) has to be measured and it is given by the expression,

$\mathbf{R = \dfrac{\mathit{l}}{\kappa\ a}}$ where κ is the conductivity of the solution.



Conductivity cell

- **Cell constant:**

i. The conductivity of an electrolytic solution is given by the expression,

$\kappa = \dfrac{1}{R}\dfrac{l}{a}$

ii. For a given conductivity cell, the quantity $\frac{l}{a}$ is a fixed quantity and is called **cell constant**.

> *The **cell constant** is defined as the ratio of the distance between the electrodes divided by the area of cross section of the electrode.*

iii. Thus, **cell constant = (G*) =** $\boldsymbol{\frac{l}{a} = \kappa . R}$

iv. Unit of cell constant: m^{-1} or cm^{-1}.

EXAMPLE – 2.8

Conductivity of a solution with 0.1 N KCl at 296 K is 0.011 ohm^{-1} cm^{-1}. Resistance of solution at same temperature is 54 ohm. Find out the cell constant of the conductivity cell.

Solution:

Conductivity, $\kappa = 0.011\ ohm^{-1}\ cm^{-1}$

Resistance, R = 54 ohm

Cell constant $(G^*) = \frac{l}{a} = \kappa \times R$

$G^* = 0.011\ ohm^{-1}\ cm^{-1} \times 54\ ohm$

$= \mathbf{0.594\ cm^{-1}}$

- **Determination of cell constant:**

i. The cell constant of a cell cannot be determined by measuring ‘*l*’ and ‘a’ because this is not only inconvenient, but also unreliable. However, it can be determined by measuring the resistance of the solution of known conductivity.

ii. Generally, a standard solution of KCl, whose conductivity at different concentrations and temperatures are known, is used for this purpose.

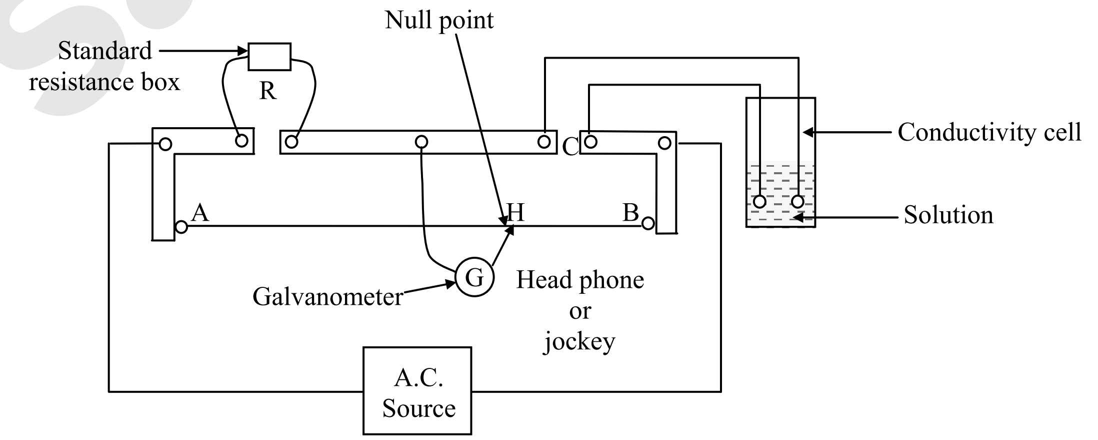
Conductivity and Molar conductivity of KCl solutions at 298.15 K:

Concentration / Molarity		**Conductivity (κ)**		**Molar Conductivity $(\wedge_m)$**	
mol L^{-1}	mol m^{-3}	S cm^{-1}	S m^{-1}	S cm^2 mol^{-1}	S m^2 mol^{-1}
1.000	1000	0.1113	11.13	111.3	111.3×10^{-4}
0.100	100.0	0.0129	1.29	129.0	129.0×10^{-4}
0.010	10.00	0.00141	0.141	141.0	141.0×10^{-4}

- **Determination of conductivity of the solution:**

i. Once the cell constant of a conductivity cell is determined, it can be used to determine the resistance or conductivity of any solution.

ii. The set up to measure the resistance of a solution (according to the Wheatstone bridge principle), is shown in the following figure.



Apparatus for resistance (or conductivity) measurement

iii. At the null point, ‘H’

$$\frac{\text{resistance C}}{\text{resistance R}} = \frac{\text{resistance of wire BH}}{\text{resistance of wire AH}} = \frac{\text{length BH}}{\text{length AH}}$$

$$\therefore \quad \text{resistance C} = \frac{\text{length BH}}{\text{length AH}} \times \text{resistance R}$$

Thus, by knowing the resistance ‘R’ and the null (or balance) point H, the resistance ‘C’ of the solution can be calculated.

iv. From the known value of resistance and cell constant, conductivity of the solution is calculated from the formula,

$$\kappa = \frac{1}{R}\frac{l}{a} \text{ or } \kappa = \frac{1}{R}.G^*$$

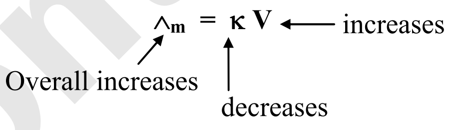
VARIATION OF CONDUCTIVITY AND MOLAR CONDUCTIVITY WITH CONCENTRATION

➢ **Variation of conductivity with concentration:**

i. The conductance of solution increases with the decrease in concentration or increase on dilution as the number of ions in solution increases on dilution.

ii. However, specific conductance or conductivity decreases upon dilution, as it is the conductance of one m^3 of the solution. Upon dilution, the concentration of ions per m^3 decreases and thereby results in decrease in conductivity. The effect of latter is more than former, so conductivity decreases with dilution.

iii. However, molar conductivity increases upon dilution, as it is the product of conductivity and volume of solution containing one mole of electrolyte. With dilution, first (former) factor decreases while the second (latter) factor increases. The increase in second factor is much more than the decrease in first factor, therefore, molar conductivity increases on dilution.



Molar conductivity increases on dilution

iv. When concentration of solution approaches zero, i.e., at infinite dilution, the molar conductivity of the solution at this dilution is known as **limiting molar conductivity** (Λ_m^o).

i.e., $\Lambda_m = \Lambda_m^o$ when $C \rightarrow 0$ (at infinite dilution)

The molar conductivity varies in a different manner for strong and weak electrolytes.

➢ **Variation of molar conductivity for strong electrolytes:**

i. Strong electrolytes ionize completely in solution at all concentrations. Their molar conductivity increases slowly with dilution (due to increase in volume).

ii. In concentrated solutions of strong electrolytes, there are strong forces of attraction between the ions of opposite charges known as interionic forces.

iii. With dilution, the ions move apart from one another with decrease in interionic forces. This results in the increase in molar conductivity of strong electrolytes with dilution.

iv. At infinite dilution, interionic attraction becomes negligible and molar conductivity reaches its limiting value.

v. Friedrich Kohlrausch showed that molar conductivity of strong electrolytes varies with concentration as:

Molar conductivity, $\boldsymbol{\Lambda_m = \Lambda_m^o - A\,C^{1/2}}$

where, Λ_m^o is the molar conductivity at zero concentration (or infinite dilution), ‘$C^{1/2}$’ is the square root of the concentration and ‘A’ is a constant.

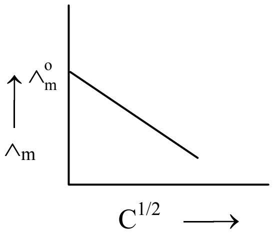
This equation is known as **Debye Huckel Onsager equation**.

This law holds good at low concentration.

vi. The graph of molar conductivity (Λ_m) versus $C^{1/2}$, for strong electrolytes is linear, especially at lower concentration. Limiting molar conductivity can be obtained by extrapolating the graph to $C^{1/2} = 0$.

vii. The slope of the line gives **‘–A’**, while intercept gives Λ_m^o.

viii. The value of ‘A’, for a given solvent and temperature, depends on the type of electrolyte, i.e., the charges on the cation and anion produced on the dissociation of the electrolyte in the solution.



Variation of Λ_m with $C^{1/2}$ for strong electrolytes

ix. Consider the following electrolytes:

Electrolyte	Charge on cation	Charge on anion	Type of electrolyte
NaCl	+1	−1	1−1
$CaCl_2$	+2	−1	2−1
$MgSO_4$	+2	−2	2−2

All the electrolytes of same type have same value of 'A'.

➢ **Variation of molar conductivity for weak electrolytes:**

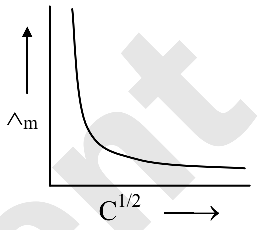
i. Weak electrolytes ionize partially in solutions and therefore, have low conductivity.

ii. Upon dilution, the value of molar conductivity for weak electrolyte increases sharply due to the increase in degree of dissociation, thereby increasing the number of ions in solution containing one mole of electrolyte.

iii. The graph of molar conductivity (Λ_m) versus $C^{1/2}$ for weak electrolytes is shown in the adjacent figure.

iv. As observed from the graph, the variation of Λ_m with $C^{1/2}$ is very large and therefore, Λ_m^o cannot be obtained by extrapolation of the graph to $C^{1/2} = 0$.

v. However, Λ_m^o for weak electrolytes can be calculated with the help of Kohlrausch's law of independent migration of ions.



Variation of Λ_m with $C^{1/2}$ for weak electrolytes

KOHLRAUSCH'S LAW AND ITS APPLICATIONS

➢ **Kohlraush's law:**

i. From a study of molar conductivities of electrolytes at infinite dilution, Kohlrausch put forward law of independent migration of ions.

> **Statement:** *The molar conductivities of an electrolyte at infinite dilution (i.e., limiting molar conductivity) is equal to the sum of molar conductivities of all the component ions (all the anions and cations) of the electrolyte.*
>
> **OR**
>
> *At infinite dilution, each ion migrates independently of its co-ion and makes its own contribution to the total molar conductivity of an electrolyte irrespective of the nature of other ion with which it is associated.*

ii. That is, $\Lambda_m^o = \lambda_+^o + \lambda_-^o$

where, Λ_m^o = molar conductivity of solution at infinite dilution,

λ_+^o = molar conductivity of cation at infinite dilution,

λ_-^o = molar conductivity of anion at infinite dilution

iii. In general, if an electrolyte on dissociation gives v_+ cations and v_- anions, then its limiting molar conductivity is given by:

$$\boldsymbol{\Lambda_m^o = v_+\lambda_+^o + v_-\lambda_-^o}$$

iv. The difference in Λ_m^o value for KCl and NaCl; KBr and NaBr and KI and NaI at 298 K is found to be 23.4 S cm^2 mol^{-1}.

i.e., $\Lambda_{m(KCl)}^o - \Lambda_{m(NaCl)}^o = \Lambda_{m(KBr)}^o - \Lambda_{m(NaBr)}^o = \Lambda_{m(KI)}^o - \Lambda_{m(NaI)}^o \simeq 23.4 \text{ S cm}^2 \text{ mol}^{-1}$

Similarly, it was found that,

$$\Lambda_{m(NaBr)}^o - \Lambda_{m(NaCl)}^o = \Lambda_{m(KBr)}^o - \Lambda_{m(KCl)}^o \simeq 1.8 \text{ S cm}^2 \text{ mol}^{-1}$$

➢ **Applications of Kohlrausch's law:**

Following are the applications of Kohlrausch's law:

- **Calculation of limiting molar conductivities $\left(\Lambda_m^o\right)$ of weak electrolytes:**

  The limiting molar conductivities, Λ_m^o, of weak electrolytes cannot be obtained by extrapolating the graph Λ_m verses $C^{1/2}$. However, this can be calculated by using Kohlrausch's law.

E.g. Limiting molar conductivity of acetic acid (CH_3COOH, a weak electrolyte) can be calculated from the known value of limiting molar conductivities of HCl, NaCl and CH_3COONa (strong electrolytes) from Kohlrausch's law as follows:

$$\wedge^{o}_{m(CH_3COONa)} = \lambda^{o}_{\left(CH_3COO^-\right)} + \lambda^{o}_{\left(Na^+\right)},\ \wedge^{o}_{m(HCl)} = \lambda^{o}_{\left(H^+\right)} + \lambda^{o}_{\left(Cl^-\right)} \text{ and } \wedge^{o}_{m(NaCl)} = \lambda^{o}_{\left(Na^+\right)} + \lambda^{o}_{\left(Cl^-\right)}$$

From this,

$$\wedge^{o}_{m(CH_3COOH)} = \lambda^{o}_{\left(CH_3COO^-\right)} + \lambda^{o}_{\left(H^+\right)} = \left[\lambda^{o}_{\left(CH_3COO^-\right)} + \lambda^{o}_{\left(Na^+\right)}\right] + \left[\lambda^{o}_{\left(H^+\right)} + \lambda^{o}_{\left(Cl^-\right)}\right] - \left[\lambda^{o}_{\left(Na^+\right)} + \lambda^{o}_{\left(Cl^-\right)}\right]$$

$$= \wedge^{o}_{m(CH_3COONa)} + \wedge^{o}_{m(HCl)} - \wedge^{o}_{m(NaCl)}$$

Limiting molar conductivity for some ions in water at 298 K:

Ion	λ° (S cm^2 mol^{-1})	Ion	λ° (S cm^2 mol^{-1})
H^+	349.6	OH^-	199.1
Na^+	50.1	Cl^-	76.3
K^+	73.5	Br^-	78.1
Ca^{2+}	119.0	CH_3COO^-	40.9
Mg^{2+}	106.0	SO_4^{2-}	160.0

- **Calculation of degree of dissociation of weak electrolyte:**
  The degree of dissociation of a weak electrolyte can be calculated from the formula,

$$\alpha = \frac{\wedge_m}{\wedge^{o}_{m}}$$

where α = degree of dissociation

$\wedge_m$ = molar conductivity of solution at concentration C

$\wedge^{o}_{m}$ = limiting molar conductivity which can be calculated from Kohlrausch's law

- **Calculation of dissociation constant of weak electrolytes:**
  The dissociation constant, K for the weak electrolyte can be calculated by the formula,

$$K = \frac{\alpha^2 C}{(1-\alpha)}$$

$$K = \frac{C\wedge^{2}_{m}}{\left(\wedge^{o}_{m}\right)^2\left(1-\frac{\wedge_m}{\wedge^{o}_{m}}\right)} \qquad \left(\text{as } \alpha = \frac{\wedge_m}{\wedge^{o}_{m}}\right)$$

$$K = \frac{C\wedge^{2}_{m}}{\wedge^{o}_{m}\left(\wedge^{o}_{m} - \wedge_m\right)}$$

Note: These formulae can also be used for calculations involving equivalent conductivity.

ELECTROLYTIC CELLS AND ELECTROLYSIS

➢ **Electrolytic cells:**

*The cells in which electric current from an external source is used to carry out non-spontaneous reactions are known as **electrolytic cells**.*

i. This process of carrying out non-spontaneous reactions in electrolytic cell is known as **electrolysis**.
ii. Therefore, **electrolysis** is a process in which an electric current is used to bring about a chemical reaction or a process in which electrical energy is converted into chemical energy.

➢ **Process of electrolysis:**

i. Two metallic rods (electrodes) are immersed in the electrolytic solution and are connected to a source of direct current (battery) with the help of metallic wires.
ii. The electrode connected to the positive terminal of the battery is called **anode** (i.e. oxidation process will take place at this electrode). While, the electrode connected to the negative terminal of the battery is called **cathode** (i.e. reduction process will take place at this electrode).

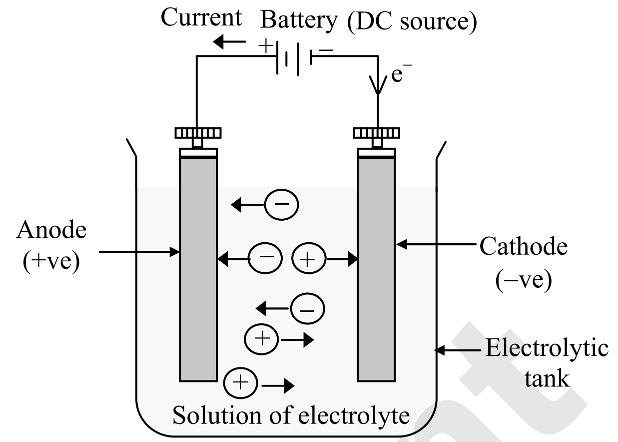
iii. When electrolysis is carried out, the ions move towards the oppositely charged electrodes.

iv. The anions on reaching the anode get oxidized (lose electrons) and cations on reaching the cathode get reduced (accept electrons).

v. Products of electrolysis depend on the nature of material being electrolysed and the type of electrodes being used.

vi. When the electrode is inert (E.g. platinum or gold) and does not participate in the chemical reaction, it acts exclusively as source or sink for electrons.

vii. Whereas, if the electrode is reactive, it participates in the electrode reaction. Thus, the products of electrolysis may be different for reactive and inert electrodes.



Process of electrolysis

Note: a. The discharge of cations and anions at their respective electrodes to form neutral species is known as **primary change**.

b. These primary products may be collected as such or may form molecules or compounds by a process known as **secondary change**.

➢ **Preferential Discharge Theory:**

i. If electrolysis of an electrolytic solution consisting of more than two ions is carried out, then certain ions are discharged in preference to others at the respective electrodes, depending upon their standard electrode potentials.

ii. At cathode, the ion having more positive value of standard electrode potential is discharged while at anode, the ion having lower value of standard electrode potential is discharged preferably.

➢ **Some Examples of Electrolysis:**

i. Electrolysis of molten NaCl:

a. Molten sodium chloride contains Na^+ and Cl^- ions, $NaCl \rightarrow Na^+ + Cl^-$

b. On passing electric current, cathode attracts Na^+ ions and anode attracts Cl^- ions.

c. Chloride ions on reaching anode lose one electron, forming chlorine atoms (primary change). These chlorine atoms being unstable, combine with other chlorine atoms, forming chlorine molecules (secondary change).

d. However, Na^+ ions on reaching cathode accept electrons forming neutral atoms.

e. The reactions may be expressed as:

At Cathode: $Na^+ + e^- \rightarrow Na$ (reduction)

At Anode: $Cl^- \rightarrow Cl + e^-$ (oxidation)

$Cl + Cl \rightarrow Cl_2$

Overall Reaction: $$\underset{\text{(molten)}}{2NaCl} \xrightarrow{\text{Electrolysis}} \underset{\text{(at cathode)}}{2Na} + \underset{\text{(at anode)}}{Cl_{2(g)}}$$

f. Therefore, the electrolysis of molten sodium chloride produces molten silvery-white sodium at cathode and pale green Cl_2 gas at anode.

ii. Electrolysis of aqueous NaCl:

a. In this case, electrolyte NaCl produces Na^+ and Cl^- ions while solvent water produces H^+ and OH^- ions as follows:

$$NaCl \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)} \quad ….(1)$$

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)} \quad ….(2)$$

b. Therefore, on passing electric current there is a competition between the reduction of Na^+ and H^+ at cathode:

$$Na^+_{(aq)} + e^- \rightarrow Na_{(s)} \quad E^o_{cell} = -2.71V….(3)$$

$$H^+_{(aq)} + e^- \rightarrow \frac{1}{2}H_{2(g)} \quad E^o_{cell} = 0.00V \quad ….(4)$$

c. As on cathode, the ion with highest reduction potential is discharged preferably; therefore, H^+ ion is discharged at cathode (formed by the dissociation of water molecule).

d. Therefore, the overall reaction is the sum of the (2) and (4) reactions and can be written as follows:

$$H_2O_{(l)} + e^- \rightarrow \frac{1}{2} H_{2(g)} + OH^-_{(aq)} \qquad ....(5)$$

e. However, at anode there is a competition between following two reactions:

$$Cl^-_{(aq)} \rightarrow \frac{1}{2} Cl_{2(g)} + e^- \qquad E^o_{cell} = 1.36 \text{ V} \quad ....(6)$$

$$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^- \qquad E^o_{cell} = 1.23\text{V} \quad ....(7)$$

f. At anode, the reaction with lower value of standard electrode potential is preferred. Therefore, the reaction involving oxidation of water having slightly less standard electrode potential should be preferred. But this reaction is found to be kinetically slow and therefore extra potential is required for the occurrence of this reaction. This extra potential is called **overpotential**. Therefore, the reaction involving oxidation of $Cl^-_{(aq)}$ is preferred (equation 6).

g. Therefore, the net reaction is: $NaCl + H_2O \rightarrow Na^+_{(aq)} + OH^-_{(aq)} + \frac{1}{2} H_{2(g)} + \frac{1}{2} Cl_{2(g)}$

h. Thus, electrolysis of aqueous sodium chloride produces hydrogen gas at cathode, chlorine gas at anode and the solution contains sodium hydroxide.

iii. Electrolysis of sulphuric acid:

a. The probable processes at anode will be:

$$2H_2O_{(l)} \longrightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^- \qquad E^o_{cell} = 1.23 \text{ V} \qquad .....(1)$$

$$2SO^{2-}_{4(aq)} \longrightarrow S_2O^{2-}_{8(aq)} + 2e^- \qquad E^o_{cell} = 1.96 \text{ V} \qquad .....(2)$$

b. For dilute sulphuric acid, reaction (1) is preferred. Whereas, at higher concentrations of H_2SO_4, reaction (2) is preferred. This can be determined on the basis of E_{cell} calculated using Nernst equation.

c. At cathode, liberation of hydrogen gas takes place as:

$$H^+_{(aq)} + e^+ \longrightarrow \frac{1}{2} H_{2(g)} \qquad E^o_{cell} = 0.00 \text{ V} \qquad ....(3)$$

FARADAY'S LAWS OF ELECTROLYSIS

In 1834, **Michael Faraday** formulated two laws on the basis of his experiments on electrolysis known as **Faraday's laws of electrolysis**.

- **Faraday's First Law of Electrolysis:**

> ***Statement:*** *The amount of chemical reaction which occurs at any electrode during electrolysis by a current is proportional to the quantity of electricity passed through the electrolyte (solution or melt).*

i. If W gram is the amount of substance deposited on passing Q coulombs of electricity, then, $W \propto Q$ and $W = ZQ$, where Z is the proportionality constant known as electrochemical equivalent.

ii. If I is the current in amperes, passed for t seconds, then, $\mathbf{W = ZIt}$

iii. If I = 1 ampere and t = 1 second then, W = Z.
Thus,

> ***Electrochemical equivalent*** *of a substance may be defined as the mass of substance deposited when a current of one ampere is passed for one second, i.e., a quantity of electricity equal to one coulomb is passed.*

iv. The amount of electricity required for the oxidation and reduction reactions, depends upon the stoichiometry of the electrode reaction.

E.g. a. One mole of silver ions require one mole of electrons for its reduction.

$$Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)}$$

b. Similarly, one mole of copper ions require two moles of electrons and one mole of aluminium ions requires three moles of electrons for their reduction as shown by the reactions:

$$Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)} \quad ; \quad Al^{3+}_{(aq)} + 3e^- \longrightarrow Al_{(s)}$$

v. The charge on one electron = 1.60218×10^{-19}C

∴ The charge on one mole of electrons = $N_A \times 1.60218 \times 10^{-19}$C

$$= 6.02214 \times 10^{23} \times 1.60218 \times 10^{-19} = 96485 \text{ C mol}^{-1} \approx \mathbf{96500\ C\ mol^{-1}}$$

This quantity of electricity is known as **Faraday** and is represented by 'F'.

vi. Therefore, one mole of silver ions (Ag^+) and one mole of copper ions (Cu^{2+}) and one mole of aluminium ions (Al^{3+}) require 1F, 2F and 3F of electricity, respectively.

EXAMPLE – 2.9

How many moles of electrons are required for reduction of 3 moles of Zn^{2+} to Zn? How many Faradays of electricity will be required?

Solution:

i. The balanced equation for the reduction of Zn^{2+} to Zn is

$Zn^{2+} + 2e^- \longrightarrow Zn_{(s)}$

The equation shows that 1 mole of Zn^{2+} is reduced to Zn by 2 moles of electrons.

For reduction of 3 moles of Zn^{2+}, **6 moles of electrons** will be required.

ii. Faraday (96500 Coulombs) is the amount of charge on one mole of electrons.

Therefore, for 6 moles of electrons, **6 F** of electricity will be required.

Alternatively, this can be represented in tabular form as given below:

Zn^{2+}	+	$2e^-$	$\longrightarrow$	Zn
1 mol Zn^{2+}		2 mol e^- = 2 F		
3 mol Zn^{2+}		6 mol e^- = 6 F		

- **Steps involved in calculating amount of substance reduced or oxidized in electrolysis:**

The mass of reactant consumed or the mass of product formed at an electrode during electrolysis can be calculated by knowing stoichiometry of the half reaction at the electrode.

i. Calculation of quantity of electricity passed: To calculate the quantity of electricity (Q) passed during electrolysis, the amount of current, I, passed through the cell is measured and the time for which the current is passed is noted.

$Q(C) = I(A) \times t(s)$

ii. Calculation of moles of electrons passed: Total charge passed is Q(C). The charge of one mole electrons is 96500 coulombs (C) or one faraday (1 F). Hence,

$$\text{Moles of electrons actually passed} = \frac{Q(C)}{96500(C / \text{mol } e^-)}$$

iii. Calculation of moles of product formed: The balanced equation for the half reaction occurring at the electrode is devised. The stoichiometry of half reaction indicates the moles of electrons passed and moles of the product formed.

From this, we will find the mole ratio, which is given by:

$$\text{Mole ratio} = \frac{\text{Moles of product formed in half reaction}}{\text{Moles of electrons required in half reaction}}$$

Moles of product formed = Moles of electrons actually passed × mole ratio

$$= \frac{Q(C)}{96500(C / \text{mol } e^-)} \times \text{mole ratio}$$

$$= \frac{I(A) \times t(s)}{96500(C / \text{mol } e^-)} \times \text{mole ratio}$$

iv. Mass of substance produced: Mass of product (W) can be calculated as given below:

W = moles of product × molar mass of product (M)

$$= \frac{I(A) \times t(s)}{96500(C / \text{mol } e^-)} \times \text{mole ratio} \times \text{molar mass of the product (M)}$$

Note: The above equation can be rearranged to calculate quantity of electricity:

$$\text{Quantity of electricity (Q)} = \frac{W \times 96500}{\text{Mole ratio} \times \text{Molar mass of product}}$$

EXAMPLE – 2.10

The mass of Cu metal produced at the cathode during the passage of 5 ampere current through $CuSO_4$ solution for 100 minutes is _______. Given: Molar mass of Cu is 63.5 g mol^{-1}.

(A) 4.935 g (B) 9.87 g

(C) 14.805 g (D) 19.74 g

Solution:

Current (I) = 5 A, t = 100 min = 100 × 60 s

Charge (Q) = 5 × 100 × 60 C

Now, the reaction for the formation of Cu is

$$Cu^{2+}_{(aq)} + 2e^{-} \longrightarrow Cu_{(s)}$$

Hence,

$$\text{Mole ratio} = \frac{\text{Moles of product formed in half reaction}}{\text{Moles of electrons required in half reaction}}$$

$$= \frac{1\ \text{mol}}{2\ \text{mol e}^{-}}$$

$$W = \frac{I(A)\times t(s)}{96500(C/\text{mol e}^{-})} \times \text{mole ratio} \times M$$

∴ Mass of Cu formed (W)

$$= \frac{5 \times 100 \times 60\ C}{96500\left(C/\text{mol e}^{-}\right)} \times \frac{1\ \text{mol}}{2\ \text{mol e}^{-}} \times 63.5\ \text{g mol}^{-1}$$

= **9.87 g**

Ans: (B)

- **Faraday's Second Law of Electrolysis:**

Statement: *The amounts of different substances liberted by the same quantity of electricity passing through the electrolytic solution are proportional to their chemical equivalent weights (Atomic Mass of Metal ÷ Number of electrons required to reduce the cation).*

Therefore, $\dfrac{\text{Weight (mass) of A deposited}}{\text{Weight (mass) of B deposited}} = \dfrac{\text{Equivalent weight of A}}{\text{Equivalent weight of B}}$

$$\frac{\text{Weight}(\text{mass})\ \text{of A deposited}}{\text{Weight}(\text{mass})\ \text{of B deposited}} = \frac{\dfrac{\text{Molar Mass of A}}{\text{Valency of A}}}{\dfrac{\text{Molar Mass of B}}{\text{Valency of B}}}$$

Since, Mole ratio = $\dfrac{1}{\text{Valency}}$

$$\frac{\text{Weight}(\text{mass})\ \text{of A deposited}}{\text{Weight}(\text{mass})\ \text{of B deposited}} = \frac{\text{Molar Mass of A} \times \text{Mole ratio of A}}{\text{Molar Mass of B} \times \text{Mole ratio of B}}$$

CAUTION

Electrochemical equivalent (Z) of a substance is not same as its equivalent weight (E).

Electrochemical equivalent (Z) is the amount of substance deposited or liberated at the electrode by the passage of 1 C of electricity through the electrolyte.

Equivalent weight (E) is the amount of substance deposited or liberated at the electrode by the passage of 1 Faraday (i.e., 96500 C) of electricity through the electrolyte.

So, $Z = \dfrac{E}{96500}$

TYPES OF CELLS OR BATTERIES

➢ **Batteries:**

i. Two or more galvanic cells connected in the series constitute **battery**. The anode of each cell is connected to the cathode of adjacent cell.

ii. Batteries can be classified in two categories:

a. Primary batteries or cells b. Secondary batteries or cells

➢ **Primary batteries or cells:**

The voltaic cells that cannot be recharged are called ***primary voltaic cells****.*

i. These types of cells once used, cannot be recharged again by applying external electric energy source. Therefore, once used, they become dead.

ii. Common example of primary cell is dry cell (Leclanche cell) and mercury cell.

➢ **Dry cell:**

i. It is the most common type of cell used in torches, toys, tape recorders, remotes and many other devices.

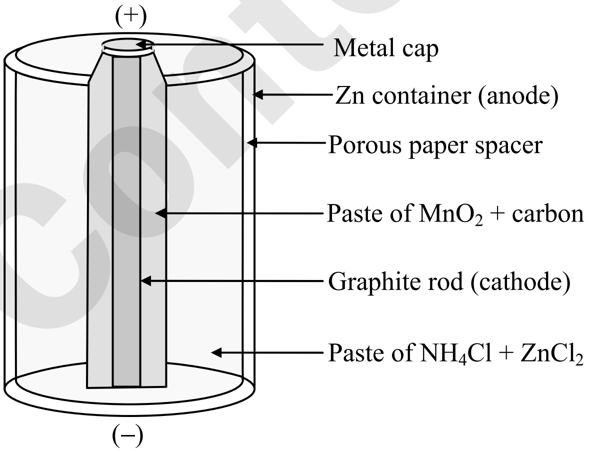
ii. It is known as dry cell as the electrolyte is a viscous aqueous paste and not a liquid solution.

iii. It is also known as **Leclanche cell** after the name of its inventor G. Leclanche.

iv. It consists of zinc container that acts as anode and a carbon (graphite) rod in the centre surrounded by powdered MnO_2 and carbon that acts as cathode.

v. The carbon rod is surrounded by black paste of manganese dioxide and carbon. Rest of the zinc container is filled with a moist paste of NH_4Cl and little of $ZnCl_2$ that acts as an electrolyte.

vi. The zinc electrode is lined from inside with a porous paper that separates it from the other materials of the cell.



Dry cell

vii. The cell is sealed at the top that prevents the drying of electrolyte paste by evaporation.

viii. When the cell operates, the zinc is oxidized to Zn^{2+}.

Anode Reaction: $Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^-$ ....(1)

ix. These electrons pass to the cathode via external circuit and cause the reduction of manganese dioxide (from electrolyte) at cathode. Therefore,

At Cathode: $MnO_2 + NH_4^+ + e^- \longrightarrow MnO(OH) + NH_3$ ....(2)

Manganese is reduced from the + 4 oxidation state to the +3 state.

Therefore, overall cell reaction is the sum of anodic reaction (1) and cathodic reaction (2).

$Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^-$ (Anodic reaction)

$MnO_2 + NH_4^+ + e^- \longrightarrow MnO(OH) + NH_3$ (Cathodic reaction)

$Zn_{(s)} + 2NH^+_{4(aq)} + 2MnO_{2(s)} \longrightarrow Zn^{2+}_{(aq)} + 2MnO(OH) + 2NH_{3(aq)}$ (Overall reaction)

x. The ammonia produced at cathode combines with zinc ions to form complex ion.

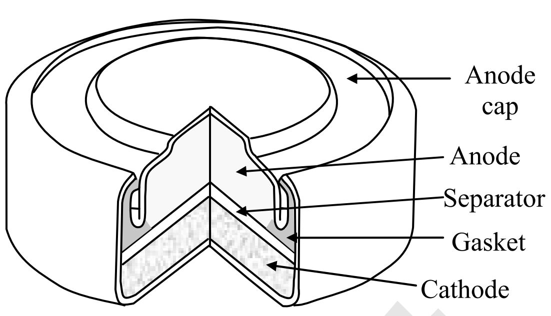
$$Zn^{2+}_{(aq)} + 4NH_{3(aq)} \longrightarrow \left[Zn\left(NH_3\right)_4\right]^{2+}_{(aq)}$$

xi. The potential of this cell is **1.5 V**.

xii. Dry cell is used in torches, flash lights, calculators, toys, tape recorders, clocks and many other electronic devices.

Note: H^+ ions from NH_4^+ result in corrosion of Zn anode, which results in the shortening of the life of dry cell.

➢ Mercury cell:



Mercury Cell

i. It is suitable for low current devices like hearing aids, watches, etc.
ii. It consists of Zn–Hg amalgam as anode and a paste of HgO and C acts as the cathode.
iii. The electrolyte is a paste made up of KOH and ZnO.
iv. The electrode reactions for the cell are given below:
At Anode: $Zn(Hg) + 2OH^{-} \longrightarrow ZnO_{(s)} + H_2O + 2e^{-}$
At Cathode: $HgO + H_2O + 2e^{-} \longrightarrow Hg_{(l)} + 2OH^{-}$
v. The overall reaction may be represented as:
$Zn(Hg) + HgO_{(s)} \longrightarrow ZnO_{(s)} + Hg_{(l)}$
vi. The cell potential remains constant during its life at approximately **1.35 V**.
vii. This is possible as the overall reaction does not involve any ion in solution whose concentration can change during its life time.

➢ Secondary batteries or cells:

The voltaic cells which can be recharged by passing electric current through it in the opposite direction and can be used again and again are called ***secondary voltaic cells****.*

i. During the recharge process, the reactions take place in opposite direction.
ii. These batteries are used in mobiles, automobiles, invertors, etc.
iii. Common examples of secondary batteries are lead storage battery and nickel-cadmium storage cell.

➢ Lead storage battery (Lead accumulators):

i. It consists of
– anode which is a series of lead grid packed with spongy lead, and
– cathode which is a series of lead grid packed with lead dioxide.
ii. These two types of plates are alternately arranged.
iii. These electrodes are immersed in an aqueous solution of 38% (by weight) of H_2SO_4 of density equal to about 1.2 g/mL, that acts as electrolyte.
iv. The cell reactions at two electrodes in the discharging state, i.e., when the cell is in use are as follows:

a. At anode, the spongy lead is oxidized to Pb^{2+} and lead plates acquire a negative charge due to the accumulation of electrons. The Pb^{2+} ions formed combine with sulphate ions (from H_2SO_4) to form insoluble $PbSO_4$.

$Pb_{(s)} \longrightarrow Pb^{2+}_{(aq)} + 2e^{-}$ (Oxidation)

$Pb^{2+}_{(aq)} + SO^{2-}_{4(aq)} \longrightarrow PbSO_{4(s)}$ (Precipitation)

$Pb_{(s)} + SO^{2-}_{4(aq)} \longrightarrow PbSO_{4(s)} + 2e^{-}$ (Overall reaction at anode)

b. The electrons produced at anode reach the cathode through external circuit. Here PbO_2 is reduced to Pb^{2+} ions in presence of H^{+} ions that combine with SO_4^{2-} ions (from H_2SO_4) to form insoluble $PbSO_4$.

$PbO_{2(s)} + 4H^{+}_{(aq)} + 2e^{-} \longrightarrow Pb^{2+}_{(aq)} + 2H_2O_{(l)}$ (Reduction)

$Pb^{2+}_{(aq)} + SO^{2-}_{4(aq)} \longrightarrow PbSO_{4(s)}$ (Precipitation)

$PbO_{2(s)} + 4H^{+}_{(aq)} + SO^{2-}_{4(aq)} + 2e^{-} \longrightarrow PbSO_{4(s)} + 2H_2O_{(l)}$ (Overall reaction at cathode)

c. Therefore, the overall cell reaction during discharge is the sum of two reactions at the two electrodes.

$Pb_{(s)} + SO^{2-}_{4(aq)} \longrightarrow PbSO_{4(s)} + 2e^{-}$ (Anodic reaction)

$PbO_{2(s)} + 4H^{+}_{(aq)} + SO^{2-}_{4(aq)} + 2e^{-} \longrightarrow PbSO_{4(s)} + 2H_2O_{(l)}$ (Cathodic reaction)

$Pb_{(s)} + PbO_{2(s)} + 2H_2SO_{4(aq)} \longrightarrow 2PbSO_{4(s)} + 2H_2O_{(l)}$ (Overall cell reaction)

v. The potential of the cell depends on concentration (density) of H_2SO_4 and during the working of cell, H_2SO_4 is consumed. Therefore, concentration of H_2SO_4 decreases and the density of the solution also decreases. The concentration of H_2SO_4 is restored; $PbSO_4(s)$ is converted to PbO_2 at anode and Pb(s) at cathode respectively, by recharging the battery.

vi. The e.m.f. of the lead storage cell is around **2 V**.

vii. The lead storage battery is used in the laboratory as a source of direct current. In automobiles, a 12 V storage battery is used by connecting six 2 V cells. It is also used in invertors.

viii. The storage battery acts as galvanic cell as well as electrolytic cell. It acts as galvanic cell during the start of the engine of the automobile as it produces electrical energy. However, it acts as an electrolytic cell during recharging.

ix. During the process of charging, the roles of anode and cathode are reversed.

i.e. PbO_2 electrode – anode – positive

and Pb electrode – cathode – negative

Thus, cell reaction during recharge,

$$PbSO_{4(s)} + 2H_2O_{(l)} \longrightarrow PbO_{2(s)} + 4H^+_{(aq)} + SO^{2-}_{4(aq)} + 2e^- \quad \text{(oxidation at anode)}$$

$$PbSO_{4(s)} + 2e^- \longrightarrow Pb_{(s)} + SO^{2-}_{4(aq)} \quad \text{(reduction at cathode)}$$

$$2PbSO_{4(s)} + 2H_2O_{(l)} \longrightarrow Pb_{(s)} + PbO_{2(s)} + 4H^+_{(aq)} + 2SO^{2-}_{4(aq)} \quad \text{(net cell reaction)}$$

Overall cell reaction during charging:

$$2PbSO_{4(s)} + 2H_2O_{(l)} \xrightarrow{\text{Charge}} Pb_{(s)} + PbO_{2(s)} + 2H_2SO_{4(aq)}$$

➢ **Nickel-cadmium (NICAD) cells**

i. It is a secondary dry cell and has longer life than lead storage cell but is more expensive.

ii. It consists of cadmium anode and nickel (IV) oxide (NiO_2) cathode supported over nickel. KOH is used as electrolyte.

iii. The reactions taking place at electrodes and overall reactions during discharge are as follows:

At Anode: $Cd_{(s)} + 2OH^-_{(aq)} \longrightarrow CdO_{(s)} + H_2O_{(l)} + 2e^-$ (Oxidation)

At Cathode: $2Ni(OH)_{3(s)} + 2e^- \longrightarrow 2Ni(OH)_{2(s)} + 2OH^-_{(aq)}$ (Reduction)

Overall Reaction: $Cd_{(s)} + 2Ni(OH)_{3(s)} \longrightarrow CdO_{(s)} + 2Ni(OH)_{2(s)} + H_2O_{(l)}$

iv. During charging of battery, reverse reaction takes place.

v. The reaction products are solid and adhere to the electrode surface and therefore can be reconverted easily during recharge of battery.

vi. The cell can be sealed as no gaseous products are formed in the reactions.

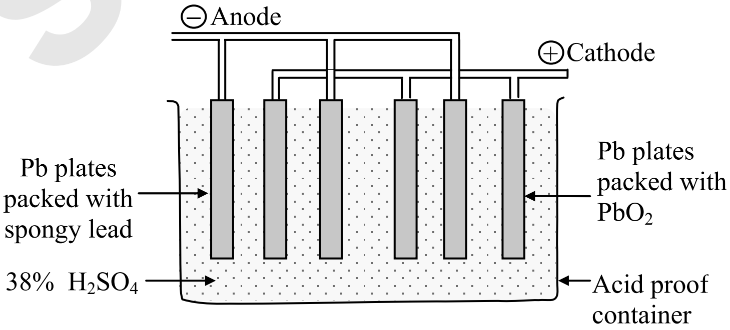
vii. It is used in calculators, electronic watches, photographic equipments, etc.

➢ **Fuel cells:**

> ***Fuel cells*** *are the galvanic cells in which the energy of combustion of fuels is directly converted into electrical energy.*

i. These fuel cells use hydrogen, methane, methanol, etc. as fuel.

ii. In these cells, the reactants are continuously fed to the electrodes and products are continuously removed from the electrolyte compartment. Therefore, these can produce electricity as long as fuel is supplied to the cell.



Lead storage battery

iii. The simplest fuel cell is **hydrogen-oxygen fuel cell**.

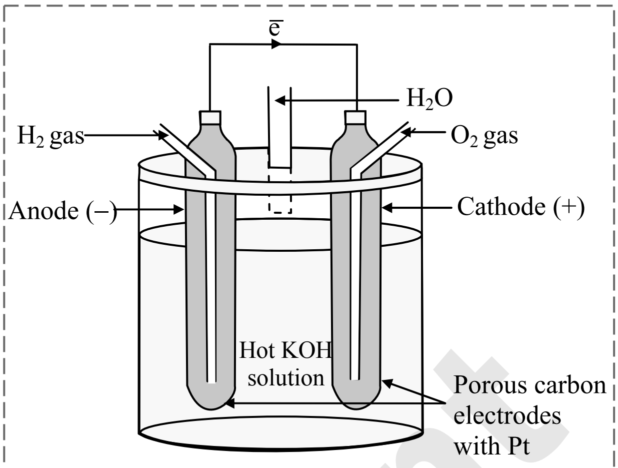
- **Hydrogen-oxygen fuel cell:**

i. Hydrogen-oxygen fuel cell consists of three compartments separated by a porous electrode. Each electrode is made up of porous compressed carbon containing a small amount of catalyst (Pt, Pd, Ag or CoO).

ii. Concentrated aqueous sodium hydroxide or potassium hydroxide is used as electrolyte and is in the central compartment.

iii. Hydrogen gas is fed into the anode compartment where it gets oxidized. Oxygen is fed into the cathode compartment where it gets reduced.

iv. The overall cell reaction produces water. The reactions can be written as:



$H_2 - O_2$ Fuel Cell

At anode: $[H_{2(g)} + 2OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + 2e^-] \times 2$ (Oxidation)

At cathode: $O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$ (Reduction)

Overall Reaction: $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

v. The potential of single cell is **1.23 V**.

vi. The first commercial use of hydrogen-oxygen fuel cell was in NASA space program (Apollo) to generate power. This cell produces water in vapour form as by product that can be condensed and used for drinking purpose. Thus, the use of these fuel cells served a dual purpose on the mission – power generation and source of pure drinking water for the astronauts.

vii. These fuel cells are pollution free and 60 – 70% efficient.

- **Advantages:**

i. During consumption of chemicals, conventional cells are discharged. However, fuel cells are not discharged because the reactants are continuously supplied.

ii. The only product of the reaction is water. Hence, fuel cells do not cause pollution.

iii. Thermal plants have efficiency of about 40% and are major source of pollution due to burning of fossil fuel. Fuel cells have an efficiency of about 70%, which is almost double.

- **Disadvantages:**

i. It can run only when reactants are continuously fed and products are continuously removed.

ii. It needs specially designed cylinder to store H_2 and O_2 in liquid form.

iii. NaOH becomes highly corrosive during combustion.

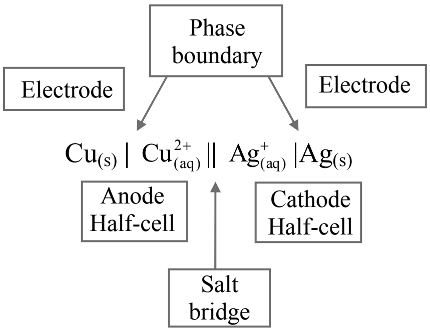
Quick Review

➢ **Electrochemical cells:**

Electrochemical cells

Galvanic or Voltaic cells	Electrolytic cells
• Conversion of chemical energy to electrical energy. • Positive electrode – Cathode – Reduction Negative electrode – Anode – Oxidation • Spontaneous chemical reaction occurs by itself and generates electricity. • **E.g.** Dry cell, fuel cell, lead storage cell, etc.	• Conversion of electrical energy to chemical energy • Positive electrode – Anode – Oxidation Negative electrode – Cathode – Reduction • Non-spontaneous reaction is forced to occur by electric current from an external source. • **E.g.** Electrorefining, electroplating, etc.

- **Cell notation:**



- **Products of electrolysis:**

Electrolysis	Products	
	Anode	**Cathode**
Molten NaCl	Chlorine gas	Molten sodium
Aqueous NaCl	Chlorine gas	Hydrogen gas
Dil. H_2SO_4	Oxygen gas	Hydrogen gas

- **Types of cells and batteries:**

Cell	Cell reaction	Uses
Dry cell or Leclanche cell: **Anode:** Zinc [container] **Cathode:** Graphite rod surrounded by black paste of MnO_2 and carbon. **Electrolyte:** A moist paste of NH_4Cl and little amount of $ZnCl_2$. **Cell potential:** 1.5 V	Anode: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ Cathode: $NH^{+}_{4(aq)} + MnO_{2(s)} + e^- \rightarrow MnO(OH)_{(s)} + NH_3$ Overall: $Zn_{(s)} + 2NH^{+}_{4(aq)} + 2MnO_{2(s)} \rightarrow Zn^{2+}_{(aq)} + 2MnO(OH)_{(s)} + 2NH_{3(aq)}$	Used in torches, toys, tape recorders, remotes and many other devices.
Mercury cell: **Anode:** Zn [in Hg] **Cathode:** A paste of HgO and carbon. **Electrolyte:** A paste of KOH and ZnO. **Cell potential:** 1.35 V	Anode : $Zn(Hg) + 2OH^- \rightarrow ZnO_{(s)} + H_2O + 2e^-$ Cathode : $HgO + H_2O + 2e^- \rightarrow Hg_{(l)} + 2OH^-$ Overall : $Zn(Hg) + HgO_{(s)} \rightarrow ZnO_{(s)} + Hg_{(l)}$	Suitable for low current devices like hearing aids, watches, etc.
Lead storage battery (Lead accumulator): **Anode:** Series of Pb grids packed with spongy Pb. **Cathode:** Series of Pb grids packed with PbO_2 **Electrolyte:** An aqueous solution of 38% (by weight) of H_2SO_4 (Density = 1.2 g/mL) The anode and cathode plates are arranged alternately.	**Process of discharging:** Anode: $Pb_{(s)} + SO^{2-}_{4(aq)} \rightarrow PbSO_{4(s)} + 2e^-$ Cathode: $PbO_{2(s)} + 4H^{+}_{(aq)} + SO^{2-}_{4(aq)} + 2e^- \rightarrow PbSO_{4(s)} + 2H_2O_{(l)}$ Overall : $Pb_{(s)} + PbO_{2(s)} + 2H_2SO_{4(aq)} \rightarrow 2PbSO_{4(s)} + 2H_2O_{(l)}$ **Process of charging:** Anode: $PbSO_{4(s)} + 2H_2O_{(l)} \rightarrow PbO_{2(s)} + 4H^{+}_{(aq)} + SO^{2-}_{4(aq)} + 2e^-$	Used in the laboratory as a source of direct current; in automobiles, a 12 V storage battery is used by connecting six 2 V cells. It is also used in invertors.

Cell potential: 2 V	Cathode: $PbSO_{4(s)} + 2e^- \rightarrow Pb_{(s)} + SO^{2-}_{4(aq)}$ Overall: $2PbSO_{4(s)} + 2H_2O_{(l)} \rightarrow Pb_{(s)} + PbO_{2(s)} + 2H_2SO_{4(aq)}$	
Nickel- cadmium (NICAD) cell: **Anode:** Cd **Cathode:** NiO_2 supported over Ni **Electrolyte:** KOH **Cell potential:** 1.4 V	**Process of discharging:** Anode: $Cd_{(s)} + 2OH^-_{(aq)} \rightarrow CdO_{(s)} + H_2O_{(l)} + 2e^-$ Cathode: $2Ni(OH)_{3(s)} + 2e^- \rightarrow 2Ni(OH)_{2(s)} + 2OH^-_{(aq)}$ Overall: $Cd_{(s)} + 2Ni(OH)_{3(s)} \rightarrow CdO_{(s)} + 2Ni(OH)_{2(s)} + H_2O_{(l)}$	Used in calculators, electronic watches, photographic equipments, etc.

- **Methods to prevent corrosion:**

Method of prevention

Painting	Galvanisation	Cathodic Protection	Passivation	Alloy formation
Coating metal surface with paint	Coating metal with zinc, which gets oxidised in preference to metal	Metal (less reactive) to be prevented from corrosion is connected to more reactive metal by wire	Metal surface made inactive towards corrosion on reaction with strong oxidising agents like HNO_3.	Homogeneous mixture of metal with other metal makes the reactive metal less suseptible to corrosion

Formulae

1. **Reduction potential** = – Oxidation Potential

2. $\mathbf{E^o_{cell}}$ **(e.m.f)** $= E°_{(cathode)} - E°_{(anode)}$

$$= E°_{(Right)} - E°_{(Left)}$$

3. **Gibbs energy change and e.m.f. of a cell:**

$\Delta_r G = -nFE_{cell}$ and $\Delta_r G° = -nFE^o_{cell}$

Where, $\Delta_r G$ = Gibbs energy change

$\Delta_r G°$ = Standard Gibbs energy change

4. **Standard cell potential and equilibrium constant:**

$$E^o_{cell} = \frac{RT}{nF}\ln K = 2.303\frac{RT}{nF}\log_{10}K$$

At 298 K, $E^o_{cell} = \dfrac{0.059}{n}\log_{10}K$

Where, K = equilibrium constant

5. **Nernst equation:**

$$E_{cell} = E^o_{cell} - \frac{2.303RT}{nF}\log_{10}\frac{[Products]}{[Reactants]}$$

$$E_{cell} = E^o_{cell} - \frac{0.059}{n}\log_{10}\frac{[Products]}{[Reactants]}$$

(At 298 K)

6. **Resistance (R):** $R = \rho\dfrac{l}{a}$

7. **Conductance (G):** $G = \dfrac{1}{R} = \dfrac{1}{\rho}\dfrac{a}{l} = \dfrac{\kappa a}{l}$

8. **Conductivity (κ):** $\kappa = \dfrac{1}{R}\dfrac{l}{a}$

9. **Molar conductivity:**

$$\wedge_m\ (S\ m^2\ mol^{-1}) = \frac{\kappa\left(Sm^{-1}\right)}{C\left(moles\ m^{-3}\right)} = \frac{G\ l\ V}{a}$$

$$\wedge_m\ (S\ cm^2\ mol^{-1}) = \frac{\kappa\left(Scm^{-1}\right)\times 1000\left(cm^3L^{-1}\right)}{C\left(mol\,L^{-1}\right)}$$

10. **Equivalent conductivity:**

$$\wedge_e\ (S\ cm^2\ eq^{-1}) = \frac{\kappa\left(Scm^{-1}\right)\times 1000\left(cm^3L^{-1}\right)}{C_{eq}\left(eq\,L^{-1}\right)}$$

Where, C_{eq} = Concentration in gram equivalent per litre (Normality)

11. **Cell constant** (G*) $= \dfrac{l}{a} = \kappa \times R$

12. Kohlrausch's Law: $\wedge_m^o = v_+\lambda_+^o + v_-\lambda_-^o$

where, $\wedge_m^o$ = Molar conductivity of solution at infinite dilution,

λ_+^o = Molar conductivity of cation at infinite dilution,

λ_-^o = Molar conductivity of anion at infinite dilution,

v_+ and v_- are no. of moles of cations and anions formed by dissociation of one mole of electrolyte.

13. Degree of dissociation (α): $\alpha = \frac{\wedge_m}{\wedge_m^o}$

14. Dissociation constant (K): $K = \frac{C\wedge_m^2}{\wedge_m^o\left(\wedge_m^o - \wedge_m\right)}$

15. Faraday's First law of electrolysis:

i. W = ZQ = ZIt

where, w = amount of substance deposited on passing Q coulombs of electricity

Z = electrochemical equivalent,

I = current in amperes, t = time in seconds

ii. Charge on 1 mole of electron = 96500 C

= 1 Faraday

Charge on n mole of electrons = nF

iii. Moles of electrons actually passed

$$= \frac{Q(C)}{96500(C/\text{mol } e^-)}$$

iv. Mole ratio

$$= \frac{\text{Moles of product formed in half reaction}}{\text{Moles of electrons required in half reaction}}$$

v. Moles of product formed

$$= \frac{I(A)\times t(s)}{96500(C/\text{mol } e^-)} \times \text{mole ratio}$$

vi. Mass of product formed (W)

$$= \frac{I(A)\times t(s)}{96500(C/\text{mol } e^-)} \times \text{mole ratio} \times \text{molar mass}$$

of the product

16. Faraday's Second law of electrolysis:

$$\frac{\text{Mass of A deposited }(W_A)}{\text{Mass of B deposited }(W_B)}$$

$$= \frac{\text{Equivalent weight of A }(E_A)}{\text{Equivalent weight of B }(E_B)}$$

$$= \frac{\text{Molar mass of A} \times \text{Mole ratio of A}}{\text{Molar mass of B} \times \text{Mole ratio of B}}$$

Multiple Choice Questions

INTRODUCTION

1. Electrochemistry deals with _______.
   (A) the production of electrical energy from chemical energy
   (B) the production of electrical energy from thermal energy
   (C) the production of electrical energy from nuclear energy
   (D) all of these

2. Electrochemistry is used _______.
   (A) in the extraction of metals like Na, K, Al, Cu
   (B) in the production of sodium hydroxide
   (C) in the extraction of non-metals like chlorine and fluorine
   (D) all of these

3. Application of electrochemistry includes _______.
   (A) electrorefining (B) electroplating
   (C) fuel cells (D) all of these

4. The study of electrochemistry is totally dependent on the concept of _______.
   (A) endothermic reaction
   (B) exothermic reaction
   (C) redox reaction
   (D) double displacement

REDOX REACTIONS

1. A redox reaction is the one in which _______.
   (A) only oxidation process takes place
   (B) only reduction process takes place
   (C) both oxidation and reduction reactions take place simultaneously
   (D) either oxidation or reduction reaction takes place

2. In following reactions, _______.
   $Na \longrightarrow Na^+ + e^-$ … (i)
   $O^{2-} - 2e^- \longrightarrow O$ … (ii)
   (A) (i) is reduction and (ii) is oxidation reaction
   (B) (i) is oxidation and (ii) is reduction reaction
   (C) both (i) and (ii) are reduction reactions
   (D) both (i) and (ii) are oxidation reactions

3. In the reaction,
   $SnCl_{2(aq)} + 2FeCl_{3(aq)} \longrightarrow SnCl_{4(aq)} + 2FeCl_{2(aq)}$
   (A) The oxidation number of Fe decreases from +3 to +2 and the oxidation number of Sn increases from +2 to +4.
   (B) The oxidation number Fe increases from +2 to +3 and the oxidation number of Sn decreases from +4 to +2.
   (C) The oxidation numbers of Fe and Sn remains same.
   (D) None of these.

4. The ______ of electrons in the reduction causes the oxidation number of an element to _______.
(A) gain, decrease (B) gain, increase
(C) loss, decrease (D) loss, increase

ELECTROCHEMICAL CELLS

1. A device in which electrical energy is produced from chemical reactions is called _______.
(A) voltaic cell
(B) galvanic cell
(C) electrochemical cell
(D) All of these

2. The chemical reaction taking place at the anode of a cell is _______.
(A) ionization (B) reduction
(C) oxidation (D) hydrolysis

3. Anode is an electrode at which, _______.
(A) electrons flow out of an electrolyte
(B) electrons flow into the electrolyte
(C) reduction takes place
(D) cations are converted to atoms

4. Which of the following reaction is possible at anode?
(A) $2Cr^{3+} + 7H_2O \longrightarrow Cr_2O_7^{2-} + 14H^+$
(B) $F_2 \longrightarrow 2F^-$
(C) $½O_2 + 2H^+ \longrightarrow H_2O$
(D) All of these

5. Which of the following is an example of an anodic reaction?
(A) $Cl + e^- \longrightarrow Cl^-$
(B) $Fe^{3+} + e^- \longrightarrow Fe^{2+}$
(C) $Zn(amalgam)+2OH^- \longrightarrow ZnO_{(s)}+H_2O+2e^-$
(D) $Na^+ + Cl^- \longrightarrow NaCl$

6. Chemical reaction taking place at cathode is _______.
(A) reduction (B) oxidation
(C) ionization (D) dissociation

7. The electrolytic decomposition of dilute sulphonic acid with platinum electrode in cathodic reaction is _______.
(A) oxidation
(B) reduction
(C) oxidation and reduction both
(D) neutralisation

8. A half-cell reaction is the one that _______.
(A) takes place at one electrode
(B) consumes half a unit of electricity
(C) goes only half way to completion
(D) involves only half a mole of electrolyte

GALVANIC OR VOLTAIC CELL

1. Which of the following is INCORRECT in a galvanic cell? **[KCET 2016]**
(A) Oxidation occurs at anode.
(B) Reduction occurs at cathode.
(C) The electrode at which electrons are gained is called cathode.
(D) The electrode at which electrons are lost is called cathode.

2. Galvanic cell converts _______.
(A) chemical energy into electrical energy
(B) electrical energy into chemical energy
(C) metal from its elemental state to the combined state
(D) electrolyte into individual ions

3. $Zn_{(s)} \mid Zn^{2+}_{(aq)}$ (1M) $\parallel Ni^{2+}_{(aq)}$ (1M) $\mid Ni_{(s)}$
Which is INCORRECT for the above given cell? **[GUJ CET 2019]**
(A) Daniel cell
(B) Galvanic cell
(C) Voltaic cell
(D) Electrochemical cell

4. Which of the following statement is TRUE for the electrochemical Daniel cell?
(A) Electrons flow from copper electrode to zinc electrode.
(B) Current flows from zinc electrode to copper electrode.
(C) Cations move towards copper electrode which is cathode.
(D) Cations move towards zinc electrode.

5. The life span of a Daniel cell is increased by _______.
(A) larger Zn electrode
(B) larger Cu electrode
(C) lowering the temperature
(D) lowering the concentration

6. The passage of current in a Daniel cell, when Cu and Zn electrodes are connected, is _______.
(A) Cu to Zn within the cell
(B) Zn to Cu within the cell
(C) Cu to Zn outside the cell
(D) Zn to Cu outside the cell

7. The initial cell potential of Daniel cell is _______.
(A) 0.5 V (B) 1.1 V
(C) 1.6 V (D) 2.1 V

8. Consider the Galvanic cell $Zn \mid ZnSO_4 \parallel CuSO_4 \mid Cu$, the reaction at cathode is _______.
(A) $Zn^{2+} + 2e^- \rightarrow Zn$
(B) $Cu^{2+} + 2e^- \rightarrow Cu$
(C) $Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$
(D) $Zn^{2+} + Cu \rightarrow Zn + Cu^{2+}$

9. An electrochemical cell can behave like an electrolytic cell when _______. **[NCERT Exemplar]**

(A) $E_{cell} = 0$ (B) $E_{cell} > E_{ext}$
(C) $E_{ext} > E_{cell}$ (D) $E_{cell} = E_{ext}$

10. In the reaction, $Cu_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$; the reduction half-cell reaction is _______.

(A) $Cu + 2e^{-} \rightarrow Cu^{2-}$ (B) $Cu - 2e^{-} \rightarrow Cu^{2+}$
(C) $Ag^{+} + e^{-} \rightarrow Ag$ (D) $Ag - e^{-} \rightarrow Ag^{+}$

11. In galvanic cell, the salt bridge is used to _______.

(A) complete the circuit
(B) reduce the electrical resistance in the cell
(C) separate cathode from anode
(D) carry salts for the chemical reaction

12. In the experiment set up for the measurement of EMF of a half cell using a reference electrode and a salt bridge, when the salt bridge is removed, the voltage _______.

(A) does not change
(B) decreases to half the value
(C) increases to maximum
(D) drops to zero

13. The salt generally used in the salt bridge of electrochemical cells is _______.

(A) KCl (B) KNO_3
(C) NH_4NO_3 (D) all of these

14. Which of the following electrolytes is used to maintain electrical neutrality in Daniel cell? **[MHT CET 2019]**

(A) KCl (B) KOH
(C) NH_4Cl (D) NaCl

15. In salt bridge, KCl is used because **[Assam CEE 2015]**

(A) KCl is present in the calomel electrode.
(B) K^{+} and Cl^{-} ions have the same transport number.
(C) K^{+} and Cl^{-} ions are isoelectronic.
(D) KCl is an electrolyte.

16. In the electrochemical cell,
$Pt \mid H_2\,(g, 1\ atm) \mid H^{+}\,(1M) \parallel Cu^{2+}\,(1M) \mid Cu(s)$
Which one of the following statements is TRUE?

(A) H_2 is cathode; Cu is anode
(B) Oxidation occurs at Cu electrode
(C) Reduction occurs at H_2 electrode
(D) H_2 is anode; Cu is cathode

17. In the cell represented by $Pb_{(s)} \mid Pb^{2+}(1\ M) \parallel Ag^{+}(1\ M) \mid Ag_{(s)}$, the reducing agent is _______. **[MHT CET 2016]**

(A) Pb (B) Pb^{2+}
(C) Ag (D) Ag^{+}

18. In the cell $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$, the negative electrode is _______.

(A) Cu (B) Cu^{2+} (C) Zn (D) Zn^{2+}

19. Which is symbolic representation for following cell reaction,
$Mg_{(s)} + Cl_{2(g)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^{-}_{(aq)}$? **[GUJ CET 2021]**

(A) $Mg \mid Mg^{2+}_{(aq)}\,(1M) \parallel Cl^{-}_{(aq)}\,(1M) \mid Cl_{2(g)}\,(1\ bar) \mid Pt$
(B) $Pt \mid Cl^{-}_{(aq)}\,(1M) \mid Cl_{2(g)}\,(1\ bar) \parallel Mg^{2+}_{(aq)}\,(1M) \mid Mg$
(C) $Mg \mid Mg^{2+}_{(aq)}\,(1M) \parallel Cl_{2(g)}\,(1\ bar) \mid Cl^{-}_{(aq)}\,(1M) \mid Pt$
(D) $Pt \mid Cl_{2(g)}\,(1\ bar) \mid Cl^{-}_{(aq)}\,(1M) \parallel Mg^{2+}_{(aq)}\,(1M) \mid Mg$

20. For the electrochemical cell, $Ag \mid AgNO_3 \parallel KCl \mid AgCl \mid Ag$; the overall cell reaction is _______.

(A) $Ag^{+} + KCl \longrightarrow AgCl_{(s)} + K^{+}$
(B) $Ag + AgCl \longrightarrow 2Ag + \frac{1}{2}Cl_2$
(C) $AgCl_{(s)} \longrightarrow Ag^{+} + Cl^{-}$
(D) $Ag^{+} + Cl^{-} \longrightarrow AgCl_{(s)}$

21. The net cell reaction for cell $Ni \mid Ni^{2+} \parallel Sn^{2+} \mid Sn$ is _______.

(A) $Sn \rightarrow Sn^{2+} + 2e^{-}$
(B) $Ni \rightarrow Ni^{2+} - 2e^{-}$
(C) $Ni^{2+} + Sn \rightarrow Ni + Sn^{2+}$
(D) $Ni + Sn^{2+} \rightarrow Ni^{2+} + Sn$

22. Silver-silver chloride electrode is a _______ electrode.

(A) metal-metal ions
(B) metal-sparingly soluble salt
(C) gas
(D) redox

23. Which of the following is a gas electrode?

(A) $H^{+}(aq) \mid H_2(g, P_{H_2}) \mid Pt$
(B) $OH^{-}\,(aq) \mid O_2(g, P_{O_2}) \mid Pt$
(C) $Cl^{-}(aq) \mid Cl_2\,(g, P_{Cl_2}) \mid Pt$
(D) All of these

24. An electrode consisting of platinum wire dipped in a solution containing the ions of the same substance in two oxidation states is a _______ electrode.

(A) metal-metal ion
(B) metal-sparingly soluble salt
(C) gas
(D) redox

25. Which of the following is an example of redox electrode?

(A) $Zn^{2+}(aq) \mid Zn$
(B) $Cl^{-}(aq) \mid AgCl(s) \mid Ag$
(C) $H^{+}(aq) \mid H_2\,(g, P_{H_2}) \mid Pt$
(D) $Fe^{2+}(aq), Fe^{3+}(aq) \mid Pt$

ELECTRODE POTENTIAL AND CELL POTENTIAL

1. The tendency of an electrode to lose electrons is known as _______.
(A) electrode potential
(B) reduction potential
(C) oxidation potential
(D) e.m.f.

2. The electrode potential depends upon _______.
(A) temperature
(B) nature of electrode
(C) concentration of ions in the solution
(D) all of these

3. In a galvanic cell, current flows _______.
(A) from anode to cathode in an external circuit
(B) from cathode to anode in an external circuit
(C) in the direction of flow of electrons
(D) from lower electrode potential to higher electrode potential

4. E.m.f. of a cell in terms of reduction potential of its left and right electrodes is _______.
(A) $E = E_{left} - E_{right}$
(B) $E = E_{left} + E_{right}$
(C) $E = E_{right} - E_{left}$
(D) $E = -(E_{Right} + E_{left})$

5. The standard oxidation potential E° for the half reactions are as:

$Zn \longrightarrow Zn^{2+} + 2e^-;\ E^o = +0.76\,V$

$Fe \longrightarrow Fe^{2+} + 2e^-;\ E^o = +0.41\,V$

The e.m.f. for cell reaction $Fe^{2+} + Zn \rightarrow Zn^{2+} + Fe$ is _______. **[WB JEE 2018]**
(A) –0.35 V (B) +0.35 V
(C) +1.2 V (D) –1.2 V

6. The standard oxidation potential of zinc and silver in water at 298 K are:

$Zn(s) \rightarrow Zn^{2+} + 2e^-;E = 0.76\,V$

$Ag(s) \rightarrow Ag^{2+} + 2e^-;E = -0.80\,V$

Which of the following reactions actually take place?
(A) $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$
(B) $Zn^{2+}(aq) + 2Ag(s) \rightarrow 2Ag^+(aq) + Zn(s)$
(C) $Zn(s) + 2Ag(s) \rightarrow Zn^{2+}(aq) + 2Ag^+(aq)$
(D) $Zn^{2+}(aq) + 2Ag^+(aq) \rightarrow Zn(s) + 2Ag(s)$

7. The cell reaction of a cell is $Mg(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Mg^{2+}(aq)$

If the standard reduction potentials of Mg and Cu are –2.37 and +0.34 V, respectively. The e.m.f. of the cell is _______.
(A) 2.03 V (B) – 2.03 V
(C) 2.71 V (D) – 2.71 V

8. Electrode potential of Zn^{2+} / Zn is –0.76 V and that of Cu^{2+}/Cu is +0.34 V. The e.m.f. of the cell constructed between these two electrodes is _______.
(A) 1.10 V (B) 0.42 V
(C) –1.1V (D) –0.42 V

9. For the electrochemical cell, $M | M^+ \| X | X^-$, $E^o(M^+/M) = 0.44$ V and $E^o(X/X^-) = 0.33$ V. From this data one can deduce that _______.
(A) $M + X \rightarrow M^+ + X^-$ is the spontaneous reaction
(B) $M^+ + X^- \rightarrow M + X$ is the spontaneous reaction
(C) $E^o_{cell} = +0.77$ V
(D) $E^o_{cell} = -0.77$ V

10. Given below are half cell reactions:

$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O,$

$E^o_{Mn^{2+}/MnO_4^-} = -1.510\ V$

$\frac{1}{2}O_2 + 2H^+ + 2e^- \longrightarrow H_2O,$

$E^o_{O_2/H_2O} = +\ 1.223\ V$

Will the permanganate ion, MnO_4^- liberate O_2 from water in the presence of an acid? **[NEET (UG) 2022]**
(A) Yes, because $E^o_{cell} = +\ 2.733$ V
(B) No, because $E^o_{cell} = -\ 2.733$ V
(C) Yes, because $E^o_{cell} = +\ 0.287$ V
(D) No, because $E^o_{cell} = -\ 0.287$ V

11. The standard electrode potentials of Zn^{2+}/Zn and Ag^+/Ag are –0.763 V and +0.799 V respectively. The standard potential of the cell is _______.
(A) 1.562 V (B) 0.036 V
(C) – 1.562 V (D) 0.799 V

12. Standard electrode potential of Ag^+ / Ag and Cu^+/Cu is $+0.80\,V$ and $+0.34\,V$, respectively. These electrodes are joint together by salt bridge if _______.
(A) copper electrode acts as cathode, then E^o_{cell} is $+0.45V$
(B) silver electrode acts as anode then E^o_{cell} is $-0.34V$
(C) copper electrode acts as anode then E^o_{cell} is $+0.46V$
(D) silver electrode acts as cathode then E^o_{cell} is $-0.34V$

13. For the cell reaction, $2Ce^{4+} + Co \rightarrow 2Ce^{3+} + Co^{2+}$; If $E^o_{Ce^{4+}/Ce^{3+}}$ is 1.89 V and $E^o_{Co/Co^{2+}}$ = – 0.28 V, calculate E^o_{cell}.

(A) – 1.64 V (B) + 1.64 V
(C) – 2.08 V (D) + 2.17 V

14. EMF of a cell whose half cells are given below is _______.

$Mg^{2+} + 2e^- \rightarrow Mg;\ E = -2.37\,V$

$Cu^{2+} + 2e^- \rightarrow Cu;\ E = +0.34\,V$

(A) – 2.03 V (B) 1.36 V
(C) 2.71 V (D) 2.03 V

15. From the following E° values of half cells;

(i) $A + e \longrightarrow A^-$; E° = – 0.24 V
(ii) $B^- + e \longrightarrow B^{2-}$; E° = + 1.25 V
(iii) $C^- + 2e \longrightarrow C^{3-}$; E° = – 1.25 V
(iv) $D + 2e \longrightarrow D^{2-}$; E° = + 0.68 V

What combination of two half cells would result in a cell with the largest potential?

(A) (ii) and (iii) (B) (ii) and (iv)
(C) (i) and (iii) (D) (i) and (iv)

16. The standard potential of Cu | Cu^{2+} electrode = –0.337 V. It corresponds to the reaction _____.

(A) $Cu \longrightarrow Cu^{2+} + 2e^-$
(B) $1/2\ Cu^{2+} + e^- \longrightarrow 1/2\ Cu$
(C) $Cu^{2+} + 2e^- \longrightarrow Cu$
(D) None of these

17. Given below are the half-cell reactions

$Mn^{2+} + 2e^- \longrightarrow Mn$; E° = –1.18 V
$Mn^{3+} + e^- \longrightarrow Mn^{2+}$; E° = + 1.51 V

The E° for $3Mn^{2+} \longrightarrow Mn + 2Mn^{3+}$ will be _______. **[JEE (Main) 2014]**

(A) –2.69 V; the reaction will not occur
(B) –2.69 V; the reaction will occur
(C) –0.33 V; the reaction will not occur
(D) –0.33 V; the reaction will occur

18. The E^o_{cell} for Zn | $Zn^{2+}_{(aq)}$ || Ni^{2+}_{aq} | Ni is _______.

$$\left(E^o_{Zn^{2+}/Zn} = -0.76\,V,\ E^o_{Ni^{2+}/Ni} = -0.25\,V\right)$$

[BCECE (Stage 1) 2016]

(A) + 0.51 V (B) + 1.1 V
(C) –1.1 V (D) –0.51 V

19. The CORRECT value of cell potential in volt for the reaction that occurs when the following two half cells are connected, is _______.

$Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$, E° = –0.44 V

$Cr_2O_7^{2-}{}_{(aq)} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$,

E° = +1.33 V

[NEET (UG) Manipur 2023]

(A) +0.01 V (B) +0.89 V
(C) +1.77 V (D) +2.65 V

20. The difference between the electrode potentials of two electrodes when no current is drawn through the cell is called _______. **[NCERT Exemplar]**

(A) cell potential
(B) cell e.m.f.
(C) potential difference
(D) cell voltage

21. E.m.f. of a cell is _______.

(A) the potential difference in an open circuit
(B) the maximum voltage delivered by a cell
(C) responsible for continuous supply of current in the cell
(D) all of these

22. Potential difference (E_{cell}) of cell is _______.

(A) less than e.m.f.
(B) more than e.m.f.
(C) equal to e.m.f.
(D) may be more or less than e.m.f.

MEASUREMENT OF ELECTRODE POTENTIAL

1. A reference electrode is used to determine the ______ of other electrode.

(A) potential (B) current
(C) resistance (D) all of these

2. Which of the following is a primary reference electrode?

(A) Calomel electrode
(B) Glass electrode
(C) Standard hydrogen electrode
(D) Silver–silver chloride electrode

3. Which of the following is a secondary reference electrode?

(A) Glass electrode
(B) Calomel electrode
(C) Silver–silver chloride electrode
(D) All of these

4. Which of the following has been universally accepted as a reference electrode at all temperatures and has been assigned a value of zero volt?

(A) Graphite electrode
(B) Copper electrode
(C) Platinum electrode
(D) Standard hydrogen electrode

5. A standard hydrogen electrode has zero electrode potential because _______.

(A) hydrogen is easiest to oxidise
(B) the electrode potential is assumed to be zero for reference
(C) hydrogen atom has only one electron
(D) hydrogen is the lightest element

6. Which one of the following is FALSE for standard hydrogen electrode?
(A) The temperature is 298 K.
(B) The concentration of H^+ ions in its solution is 1 M.
(C) The hydrogen gas pressure is two bar.
(D) An electrode of Pt, coated with platinum black is used.

7. Which of the following statement is CORRECT about an inert electrode in a cell? **[NCERT Exemplar]**
(A) It does not participate in the cell reaction.
(B) It provides surface either for oxidation or for reduction reaction.
(C) It provides surface for conduction of electrons.
(D) All of these.

8. Which cell will measure standard electrode potential of copper electrode? **[NCERT Exemplar]**
(A) Pt (s) | H_2 (g, 0.1 bar) | H^+ (aq., 1 M) || Cu^{2+} (aq., 1 M) | Cu
(B) Pt (s) | H_2 (g, 1 bar) | H^+ (aq.,1 M) || Cu^{2+} (aq., 2 M) | Cu
(C) Pt (s) | H_2 (g, 1 bar) | H^+ (aq.,1 M) || Cu^{2+} (aq., 1 M) | Cu
(D) Pt (s) | H_2 (g, 1 bar) | H^+ (aq., 0.1 M) || Cu^{2+} (aq., 1 M) | Cu

9. Which of the following is INCORRECT with respect to the difficulty involved in setting up SHE?
(A) Pure and dry H_2 gas cannot be easily obtained.
(B) It is difficult to maintain pressure of H_2 gas at exact 1 bar.
(C) It is difficult to obtain platinised platinum foil.
(D) It is difficult to maintain the concentration of H^+ ion solution at 1 M.

10. A calomel electrode is used as reference electrode because _______.
(A) its potential is arbitrarily fixed
(B) its potential is exactly known with respect to a standard
(C) the effects of temperature on e.m.f. of electrode is negligible
(D) both (B) and (C)

11. Calomel is _______. **[BCECE 2014]**
(A) $HgCl_2$ (B) Hg_2Cl_2
(C) HgI_2 (D) HgO

12. The electrode potential of calomel electrode, used as a reference electrode _______.
(A) is taken as zero always
(B) is taken as zero if concentration of KCl solution is 1 M
(C) can be zero, negative or positive
(D) depends upon the concentration of KCl solution used

13. The calomel electrode can be represented as _______.
(A) KCl (sat.) | Hg(*l*) | $HgCl_2$(s)
(B) KCl (sat.) | $HgCl_2$(s) | Hg(*l*)
(C) KCl (sat.) | Hg_2Cl_2(s) | Hg(*l*)
(D) KCl (sat.) | $AgCl_2$(s) | Hg(*l*)

14. A decinormal calomel electrode contains _______.
(A) N/10 solution of Hg_2Cl_2
(B) 1 N solution of KCl
(C) 1 N solution of Hg_2Cl_2
(D) N/10 solution of KCl

ELECTROCHEMICAL SERIES (ELECTROMOTIVE SERIES)

1. The reduction potential is the tendency of an electrode to _______.
(A) get reduced
(B) lose electrons
(C) gain electrons
(D) both (A) and (C)

2. When a half reaction is reversed, _______.
(A) the sign of E° changes, but its magnitude remains the same
(B) the magnitude of E° changes, but its sign remains the same
(C) the sign and the magnitude of E° change
(D) the sign and the magnitude of E° remain same

3. When the half reaction is multiplied by a numerical factor n, the magnitude of E° _______.
(A) should be multiplied by n
(B) should be multiplied by 2n
(C) should be multiplied by $\frac{n}{2}$
(D) remains unaffected

4. In relation to electrochemical cells, stronger the oxidizing agent, greater is the _______.
(A) oxidation potential
(B) reduction potential
(C) ionic behaviour
(D) none of these

5. If the half-cell reaction $A + e^- \to A^-$ has a large negative reduction potential, it follows that _______.
(A) A is readily reduced
(B) A is readily oxidized
(C) A^- is readily reduced
(D) A^- is readily oxidized

6. E° for $Fe^{2+} + 2e^- \longrightarrow Fe$ is –0.44 volts and E° for $Zn^{2+} + 2e^- \longrightarrow Zn$ is –0.76 volts. It means that _______.
(A) Fe is more electropositive
(B) Zn is more electropositive
(C) Zn is more electronegative
(D) all are incorrect

7. Reduction potential of four elements P, Q, R, S is – 2.90 V, + 0.34 V, + 1.20 V and – 0.76 V respectively. Reactivity decreases in the order, _______.
(A) $P > Q > R > S$ (B) $Q > P > R > S$
(C) $R > Q > S > P$ (D) $P > S > Q > R$

8. Four alkali metals A, B, C and D are having respectively standard electrode potentials as –3.05, –1.66, –0.40 and 0.80 V. Which one will be the most reactive?
(A) A (B) B
(C) C (D) D

9. K, Ca and Li metals may be arranged in the decreasing order of their standard electrode potentials as _______.
(A) K, Ca, Li (B) Ca, K, Li
(C) Li, Ca, K (D) Ca, Li, K

10. Arrange Mg, K, Ba and Ca in the order of their decreasing electrode potential.
(A) K, Ba, Ca, Mg (B) Ca, Mg, K, Ba
(C) Ba, Ca, K, Mg (D) Mg, Ca, Ba, K

11. Which of the following has the highest electrode potential?
(A) Li (B) Cu (C) Au (D) Al

12. Identify the reaction from following having top position in EMF series (Std. red. potential) according to their electrode potential at 298 K.
[NEET (UG) P-II 2020]
(A) $K^+ + 1e^- \rightarrow K_{(s)}$
(B) $Mg^{2+} + 2e^- \rightarrow Mg_{(s)}$
(C) $Fe^{2+} + 2e^- \rightarrow Fe_{(s)}$
(D) $Au^{3+} + 3e^- \rightarrow Au_{(s)}$

13. A cell constructed by coupling a standard copper electrode and a standard magnesium electrode, has e.m.f. of 2.7 volts. If the standard reduction potential of copper electrode is + 0.34 volt; that of magnesium electrode is _______.
(A) + 3.04 volts (B) – 3.04 volts
(C) + 2.36 volts (D) – 2.36 volts

14. If the reduction potential of a species is more, then _______.
(A) it is easily oxidized
(B) it acts as a reducing agent
(C) it acts as an oxidizing agent
(D) it has redox nature

15. Consider the following standard electrode potentials (E° in volts) in aqueous solution:

Element	M^{3+}/M	M^+/M
Al	-1.66	+0.55
Tl	+1.26	-0.34

Based on these data, which of the following statements is CORRECT?
(A) Tl^+ is more stable than Al^{3+}
(B) Al^+ is more stable than Al^{3+}
(C) Tl^+ is more stable than Al^+
(D) Tl^{3+} is more stable than Al^{3+}

16. The value of E^o_{red} for metals A, B and C are 0.34 volt, –0.80 volt and –0.46 volt respectively. State the CORRECT order for their ability to act as reducing agent. **[GUJ CET 2015]**
(A) C > B > A (B) A > B > C
(C) B > C > A (D) C > A > B

17. Identify the weakest oxidizing agent among the following. **[MHT CET 2017]**
(A) Li^+ (B) Na^+
(C) Cd^{2+} (D) I_2

18. Given that
$E^o_{O_2/H_2O} = +1.23$ V; $E^o_{S_2O_8^{2-}/SO_4^{2-}} = 2.05$ V;
$E^o_{Br_2/Br^-} = +1.09$ V; $E^o_{Au^{3+}/Au} = +1.4$ V
the strongest oxidizing agent is _______.
[JEE (Main) April 2019]
(A) $S_2O_8^{2-}$ (B) O_2
(C) Br_2 (D) Au^{3+}

19. Using the data given below, find out the strongest reducing agent.
$E^o_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33\,V; E^o_{Cl_2/Cl^-} = 1.36\,V$
$E^o_{MnO_4^-/Mn^{2+}} = 1.51\,V; E^o_{Cr^{3+}/Cr} = -0.74\,V$
[NCERT Exemplar; JEE (Main) 2017; GUJ CET 2021]
(A) Cl^- (B) Cr (C) Cr^{3+} (D) Mn^{2+}

20. The standard reduction potentials at 25 °C of Li^+/Li, Ba^{2+}/Ba, Na^+/Na and Mg^{2+}/Mg are –3.05, –2.73, –2.71 and –2.37 volts, respectively. Which one of the following is the strongest oxidizing agent ?
(A) Na^+ (B) Li^+ (C) Ba^{2+} (D) Mg^{2+}

21. Use the data given and find out the most stable ion in its reduced form.
$E^o_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33\,V; E^o_{Cl_2/Cl^-} = 1.36\,V$
$E^o_{MnO_4^-/Mn^{2+}} = 1.51\,V; E^o_{Cr^{3+}/Cr} = -0.74\,V$
[NCERT Exemplar]
(A) Cl^- (B) Cr^{3+} (C) Cr (D) Mn^{2+}

22. Electrode potential data are given as:

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq);\ E^o = +0.77\ V$

$Al^{3+}(aq) + 3e^- \rightarrow Al(s);\ E^o = -1.66\ V$

$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq);\ E^o = +1.08\ V$

Based on the data given above, reducing power of Fe^{2+}, Al and Br^- will increase in the order _______.

(A) $Br^- < Fe^{2+} < Al$ (B) $Fe^{2+} < Al < Br^-$
(C) $Al < Br^- < Fe^{2+}$ (D) $Al < Fe^{2+} < Br^-$

23. Using the data given, find out in which option the order of reducing power is CORRECT.

$E^o_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33\,V;\ E^o_{Cl_2/Cl^-} = 1.36\,V$

$E^o_{MnO_4^-/Mn^{2+}} = 1.51\,V;\ E^o_{Cr^{3+}/Cr} = -0.74\,V$

[NCERT Exemplar]

(A) $Cr^{3+} < Cl^- < Mn^{2+} < Cr$
(B) $Mn^{2+} < Cl^- < Cr^{3+} < Cr$
(C) $Cr^{3+} < Cl^- < Cr_2O_7^{2-} < MnO_4^-$
(D) $Mn^{2+} < Cr^{3+} < Cl^- < Cr$

24. The standard electrode potential (E°) values of Al^{3+}/Al, Ag^+/Ag, K^+/K and Cr^{3+}/Cr are –1.66 V, 0.80 V, –2.93 V and –0.74 V, respectively. The CORRECT decreasing order of reducing power of the metal is _______.

[NEET (UG) Odisha 2019]

(A) Al > K > Ag > Cr
(B) Ag > Cr > Al > K
(C) K > Al > Cr > Ag
(D) K > Al > Ag > Cr

25. The standard reduction potential at 298 K for the following half-cell reactions are,

$Zn^{2+}_{(aq)} + 2e^- \longrightarrow Zn_{(s)} \quad E° = -0.762\ V$

$Cr^{3+}_{(aq)} + 3e^- \longrightarrow Cr_{(s)} \quad E° = -0.740\ V$

$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)} \quad E° = 0.0\ V$

$F_{2(g)} + 2e^- \longrightarrow 2F^-_{(aq)} \quad E° = 2.87\ V$

Which of the following is the strongest reducing agent? **[KCET 2017]**

(A) $Zn_{(s)}$ (B) $Cr_{(s)}$ (C) $H_{2(g)}$ (D) $F_{2(g)}$

26. Which one among the following is the strongest reducing agent?

$Fe^{2+} + 2e^- \rightarrow Fe(-0.44\,V)$

$Ni^{2+} + 2e^- \rightarrow Ni(-0.25\,V)$

$Sn^{2+} + 2e^- \rightarrow Sn(-0.14\,V)$

$Fe^{3+} + e^- \rightarrow Fe^{2+}(+0.77\,V)$

(A) Fe (B) Fe^{2+} (C) Ni (D) Sn

27. The standard reduction potentials of 4 elements are given below. Which of the following will be the most suitable reducing agent?

I = – 3.04 V, II = – 1.90 V, III = 0 V,
IV = 1.90 V

(A) I (B) II (C) III (D) IV

28. E° values of Mg^{2+}/Mg is –2.37 V, of Zn^{2+}/Zn is –0.76 V and Fe^{2+}/Fe is –0.44 V. Which of the following statements is CORRECT?

(A) Zn will reduce Fe^{2+}
(B) Zn will reduce Mg^{2+}
(C) Mg oxidizes Fe
(D) Zn oxidizes Fe

29. When $E^o_{Ag^+/Ag} = 0.80$ volt and $E^o_{Zn^{2+}/Zn} = -0.76$ volt, which of the following is CORRECT?

(A) Ag^+ can be reduced by H_2
(B) Ag can be oxidized by H_2
(C) Zn^{2+} can be reduced by H_2
(D) Ag can reduce Zn^{2+} ion

30. Zinc displaces copper from the solution of its salt because _______.

(A) atomic number of zinc is more than that of copper
(B) zinc salt is more soluble in water than the copper salt
(C) Gibbs free energy of zinc is less than that of copper
(D) standard electrode potential of zinc is lower than copper

31. Standard reduction potential of Cu and Zn is + 0.34 V and –0.763 V, respectively. Using the above data, choose the CORRECT option?

(A) $CuSO_4$ solution can be stored in Zn vessel.
(B) $ZnSO_4$ solution can be stored in Cu vessel.
(C) With respect to Cu, Zn is a stronger oxidizing agent.
(D) Cu as well as Zn will displace hydrogen from dilute HCl.

32. In which metal container, the aqueous solution of $CuSO_4$ can be stored? **[GUJ CET 2018]**

$E^o_{Cu^{2+}/Cu} = 0.34\ V,\quad E^o_{Fe/Fe^{2+}} = 0.44\ V,$

$E^o_{Al/Al^{3+}} = 1.66\ V,\quad E^o_{Ni/Ni^{2+}} = 0.25\ V,$

$E^o_{Ag^+/Ag} = 0.80\ V$

(A) Ag (B) Ni (C) Fe (D) Al

33. If an iron rod is dipped in $CuSO_4$ solution, _______.

(A) blue colour of the solution turns green
(B) brown layer is deposited on iron rod
(C) no change occurs in the colour of the solution
(D) both (A) and (B)

34. The pair of compounds that can exist together is ________. **[AIPMT 2014]**

(A) $FeCl_3$, $SnCl_2$ (B) $HgCl_2$, $SnCl_2$
(C) $FeCl_2$, $SnCl_2$ (D) $FeCl_3$, KI

35. The standard potential at 25 °C for the following half reactions are given:

$Zn^{2+} + 2e \rightarrow Zn, E^{\circ} = -0.762\,V$;

$Mg^{2+} + 2e \rightarrow Mg, E^{\circ} = -2.37\,V$

When zinc dust is added to the solution of $MgCl_2$, _______

(A) $ZnCl_2$ is formed
(B) Zinc dissolves in the solution
(C) no reaction takes place
(D) Mg is precipitated

36. Which of the following metal does NOT react with the solution of copper sulphate?

(A) Mg (B) Fe (C) Zn (D) Ag

37. On the basis of position in the electrochemical series, the metal, which does NOT displace hydrogen from water and acids, is _______.

(A) Hg (B) Al (C) Pb (D) Ba

38. The position of some metals in the electrochemical series in decreasing electropositive character is given as $Mg > Al > Zn > Cu > Ag$. What will happen, if a copper spoon is used to stir a solution of aluminium nitrate?

(A) The spoon will get coated with Al.
(B) An alloy of Cu and Al is formed.
(C) The solution will become blue.
(D) There will be no reaction.

39. When the sample of copper with zinc impurity is to be purified by electrolysis, the appropriate electrodes are:

	Cathode	**Anode**
(A)	Pure zinc	Pure copper
(B)	Impure sample	Pure copper
(C)	Impure zinc	Impure sample
(D)	Pure copper	Impure sample

40. In the process of electroplating of an object with nickel, _______.

(A) the object and the pure nickel rod are made the anode.
(B) pure nickel rod and the object are made the cathode.
(C) the object is made the cathode and a pure nickel rod, the anode.
(D) the object is made the anode and a pure nickel rod, the cathode.

41. In which of the following cases a chemical reaction is possible? **[KCET 2020]**

(A) $AgNO_3$ solution is stirred with a copper spoon.
(B) Conc. HNO_3 is stored in a platinum vessel.
(C) Gold ornaments arc washed with dil HCl.
(D) $ZnSO_{4(aq)}$ is placed in a copper vessel.

42. The metal that CANNOT be obtained by electrolysis of an aqueous solution of its salt is _______. **[JEE (Main) 2014]**

(A) Ag (B) Ca
(C) Cu (D) Cr

43. At 298 K, the standard electrode potentials of Cu^{2+}/Cu, Zn^{2+}/Zn, Fe^{2+}/Fe and Ag^{+}/Ag are 0.34 V, –0.76 V, –0.44 V and 0.80 V, respectively.

On the basis of standard electrode potential, predict which of the following reaction CANNOT occur? **[NEET (UG) 2022]**

(A) $FeSO_{4(aq)} + Zn_{(s)} \longrightarrow ZnSO_{4(aq)} + Fe_{(s)}$
(B) $2CuSO_{4(aq)} + 2Ag_{(s)} \longrightarrow 2Cu_{(s)} + Ag_2SO_{4(aq)}$
(C) $CuSO_{4(aq)} + Zn_{(s)} \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$
(D) $CuSO_{4(aq)} + Fe_{(s)} \longrightarrow FeSO_{4(aq)} + Cu_{(s)}$

44. The standard oxidation potential of Zn and Ag at 20°C are:

$Zn_{(s)} \longrightarrow Zn^{2+}{}_{(aq)} + 2e^{-}$; $E^{\circ} = 0.76$ V

$Ag_{(s)} \longrightarrow Ag^{+}{}_{(aq)} + e^{-}$; $E^{\circ} = -0.80$ V

Which of the following reactions actually takes place?

(A) $Zn_{(s)} + 2Ag^{+}{}_{(aq)} \longrightarrow Zn^{2+}{}_{(aq)} + 2Ag_{(s)}$
(B) $Zn^{2+}{}_{(aq)} + 2Ag_{(s)} \longrightarrow 2Ag^{+}{}_{(aq)} + Zn_{(s)}$
(C) $Zn_{(s)} + Ag_{(s)} \longrightarrow Zn^{2+}{}_{(aq)} + Ag^{+}{}_{(aq)}$
(D) $Zn^{2+}{}_{(aq)} + Ag^{+}{}_{(aq)} \longrightarrow Zn_{(s)} + Ag_{(s)}$

45. Consider the change in oxidation state of bromine corresponding to different emf values as shown in the diagram below:

$$BrO_4^{-} \xrightarrow{1.82\ V} BrO_3^{-} \xrightarrow{1.5\ V} HBrO \xrightarrow{1.595\ V} Br_2 \xrightarrow{1.0652\ V} Br^{-}$$

Then the species undergoing disproportionation is _______. **[NEET (UG) 2018]**

(A) BrO_3^{-} (B) BrO_4^{-}
(C) Br_2 (D) HBrO

RELATION BETWEEN GIBBS ENERGY CHANGE AND E.M.F. OF A CELL

1. The electrical work done in a galvanic cell is equal to _____.

(A) E_{cell} (B) FE_{cell}
(C) nFE_{cell} (D) $\frac{E_{cell}}{nF}$

2. Which of the following expression is CORRECT?

(A) $\Delta G^{\circ} = -nF\,E^{o}_{cell}$
(B) $\Delta G^{\circ} = nF\,E^{o}_{cell}$
(C) $\Delta G^{\circ} = -2.303RT\,nFE^{o}_{cell}$
(D) $\Delta G^{o} = 2.303\,RT\log_{10}K$

3. Given below are two statements: one is labelled as **Assertion A** and the other is labelled as **Reason R**.

Assertion A: In equation $\Delta_rG = -nFE_{cell}$, value of Δ_rG depends on n.

Reason R: E_{cell} is an intensive property and Δ_rG is an extensive property.

In the light of the above statements, choose the CORRECT answer from the options given below: **[NEET (UG) 2023]**

(A) Both **A** and **R** are true and **R** is NOT the correct explanation of **A**.
(B) **A** is true but **R** is false.
(C) **A** is false but **R** is true.
(D) Both **A** and **R** are true and **R** is the correct explanation of **A**.

4. Which of the following statements is CORRECT? **[NCERT Exemplar]**

(A) E_{cell} and Δ_rG of cell reaction both are extensive properties.
(B) E_{cell} and Δ_rG of cell reaction both are intensive properties.
(C) E_{cell} is an intensive property while Δ_rG of cell reaction is an extensive property.
(D) E_{cell} is an extensive property while Δ_rG of cell reaction is an intensive property.

5. The standard cell potential of:
Zn/Zn^{2+} (aq) || Cu^{2+} (aq) | Cu cell is 1.10 V.
The maximum work obtained by this cell will be _______.

(A) –106.15 kJ (B) – 212.30 kJ
(C) – 318.45 kJ (D) – 424.60 kJ

6. If the ΔG° of a cell reaction, $AgCl + e^- \rightarrow Ag + Cl^-$ is -21.20 kJ, the standard e.m.f. of cell is _______.

(A) 0.229 V (B) 0.220 V
(C) – 0.220 V (D) – 0.110 V

7. For the cell reaction,
$2Fe^{3+}_{(aq)} + 2I^-_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(aq)}$
E^o_{cell} = 0.24 V at 298 K. The standard Gibbs energy (Δ_rG^o) of the cell reaction is _______.
[Given that Faraday constant F = 96500 C mol^{-1}] **[NEET (UG) 2019]**

(A) –23.16 kJ mol^{-1}
(B) 46.32 kJ mol^{-1}
(C) 23.16 kJ mol^{-1}
(D) –46.32 kJ mol^{-1}

8. What is the free energy change per mole of Cu(II) ion formed in a cell consisting of Cu|Cu(II)ion half-cell suitably connected to a Ag|Ag^+ ion half-cell? (Given: E° = 0.46 V) **[Assam CEE 2015]**

(A) –75 kJ (B) –89 kJ
(C) –45 kJ (D) –25 kJ

9. What is the standard reduction potential (E°) for $Fe^{3+} \rightarrow Fe$?
Given that :

$$Fe^{2+} + 2e^- \longrightarrow Fe;\ E^o_{Fe^{2+}/Fe} = -0.47\ V$$

$$Fe^{3+} + e^- \longrightarrow Fe^{2+};\ E^o_{Fe^{3+}/Fe^{2+}} = +0.77\ V$$

(A) –0.057 V (B) +0.057 V
(C) +0.30 V (D) –0.30V

10. Given: $E^o_{Mn^{7+}/Mn^{2+}}$ = 1.5 V and $E^o_{Mn^{4+}/Mn^{2+}}$ = 1.2 V, then $E^o_{Mn^{7+}/Mn^{4+}}$ is _______. **[KCET 2019]**

(A) 1.7 V (B) 2.1 V
(C) 0.3 V (D) 0.1 V

11. Given that the standard potentials (E°) of Cu^{2+}/Cu and Cu^+/Cu are 0.34 V and 0.522 V respectively, the E° of Cu^{2+}/Cu^+ is _______. **[JEE (Main) Jan 2020]**

(A) +0.158 V (B) –0.182 V
(C) 0.182 V (D) –0.158 V

12. Calculate the standard cell potential (in volt) of the cell in which the following reaction takes place:
$Fe^{2+}_{(aq)} + Ag^+_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + Ag_{(s)}$
Given that:
$E^o_{Ag^+/Ag}$ = X V ; $E^o_{Fe^{2+}/Fe}$ = Y V
$E^o_{Fe^{3+}/Fe}$ = Z V

(A) X + 2Y – 3Z (B) X + Y – 3Z
(C) X + Y + Z (D) X + 2Y – Z

13. For hydrogen–oxygen fuel cell at one atm and 298 K, $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)}$; ΔG° = –240 kJ
E° for the cell is approximately _______.
(Given 1F = 96500 C mol^{-1}) **[KCET 2014]**

(A) 2.48 V (B) 1.24 V
(C) 2.5 V (D) 1.26 V

14. For a spontaneous reaction, the Gibb's free energy change (ΔG°), equilibrium constant (K) and cell potential (E^o_{Cell}) will be respectively _______.

(A) –ve, > 1, +ve (B) +ve, > 1, –ve
(C) –ve, < 1, –ve (D) –ve, > 1, –ve

15. If the E^o_{cell} for a given reaction has a negative value, which of the following gives the CORRECT relationships for the values of ΔG° and K_{eq}? **[NEET (UG) P-II 2016]**

(A) ΔG° < 0; K_{eq} < 1
(B) ΔG° > 0; K_{eq} < 1
(C) ΔG° > 0; K_{eq} > 1
(D) ΔG° < 0; K_{eq} > 1

16. The relationship between standard reduction potential of cell and equilibrium constant at 298 K is shown by _______.

(A) $E^o_{cell} = \frac{n}{0.0591}\log_{10} K$

(B) $E^o_{cell} = \frac{0.0591}{n}\log_{10} K$

(C) $E^o_{cell} = 0.0591\, n \log_{10} K$

(D) $E^o_{cell} = \frac{\log_{10} K}{n}$

17. E^o_{cell} = 1.1 V for Daniel cell. Which of the following expressions are CORRECT description of state of equilibrium in this cell? **[NCERT Exemplar]**

i. $1.1 = K_C$

ii. $\frac{2.303RT}{2F}\log K_C = 1.1$

iii. $\log K_C = \frac{2.2}{0.059}$

iv. $\log K_C = 1.1$

(A) i, ii
(B) ii, iii
(C) i, iii
(D) ii, iv

18. For a cell involving one electron E^o_{cell} = 0.59 V at 298 K, the equilibrium constant for the cell reaction is _______.

[Given that $\frac{2.303\ RT}{F} = 0.059$ V at T = 298 K]

[NEET (UG) 2019]

(A) 1.0×10^5
(B) 1.0×10^{10}
(C) 1.0×10^{30}
(D) 1.0×10^2

19. E° for the cell $Zn \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu$ is 1.10 V at 25 °C, the equilibrium constant for the reaction $Zn + Cu^{2+}(aq) \rightleftharpoons Cu + Zn^{2+}(aq)$ is of the order of _______.

(A) 10^{-28}
(B) 10^{+37}
(C) 10^{-18}
(D) 10^{+17}

20. Given $E^o_{Fe^{+3}/Fe^{+2}} = +0.76$ V and $E^o_{I_2/I^-} = +0.55$ V. The equilibrium constant for the reaction taking place in galvanic cell consisting of above two electrodes is _______.

$\left[\frac{2.303\ RT}{F} = 0.06\right]$ **[KCET 2020]**

(A) 1×10^9
(B) 1×10^8
(C) 5×10^{12}
(D) 1×10^7

NERNST EQUATION AND ITS APPLICATIONS

1. E_{cell} of a cell; $aA + bB \longrightarrow cC + dD$ is _______.

(A) $E^o_{cell} - \frac{RT}{nF}\ln\frac{[C]^c[D]^d}{[A]^a[B]^b}$

(B) $E^o_{cell} - RT\ln\frac{[a]^A[b]^B}{[c]^C[d]^D}$

(C) $E^o_{cell} - \frac{RT}{nF}\ln\frac{[C]^c[d]^D}{[A]^a[B]^b}$

(D) $E^o_{cell} - \frac{RT}{nF}\ln\frac{[C]^c[d]^D}{[a]^A[B]^b}$

2. Which of the following is CORRECT expression for electrode potential of a cell?

(A) $E = E^o - \frac{RT}{nF}\ln\frac{[products]}{[reactants]}$

(B) $E = E^o + \frac{RT}{F}\ln\frac{[products]}{[reactants]}$

(C) $E = E^o - \frac{RT}{nF}\ln\frac{[reactants]}{[products]}$

(D) $E = -\frac{RT}{F}\ln\frac{[products]}{[reactants]}$

3. In Nernst equation, the constant 0.0591 at 298 K represents _______.

(A) $\frac{2.303RT}{nF}$

(B) $\frac{2.303RT}{F}$

(C) $\frac{RT}{nF}$

(D) $\frac{2.303RT}{nF}\log_{10}\frac{\text{reduced state}}{\text{oxidised state}}$

4. The CORRECT representation of Nernst's equation at 298 K is _______.

(A) $E_{M^{n+}/M} = E^o_{M^{n+}/M} + \frac{0.059}{n}\log[M^{n+}]$

(B) $E_{M^{n+}/M} = E^o_{M^{n+}/M} - \frac{0.059}{n}\log[M^{n+}]$

(C) $E_{M^{n+}/M} = E^o_{M^{n+}/M} + \frac{n}{0.059}\log[M^{n+}]$

(D) All of these

5. $Zn(s) + Cl_2(1\ atm) \rightarrow Zn^{2+} + 2Cl^-$. E^o_{cell} of the cell is 2.12 V. To increase E_{cell}, _______.

(A) $[Zn^{2+}]$ should be increased
(B) $[Zn^{2+}]$ should be decreased
(C) $[Cl^-]$ should be decreased
(D) P_{Cl_2} should be decreased

6. At temperature of 298 K, the e.m.f. of the following electrochemical cell $Ag_{(s)} \mid Ag^{+}$ (0.1 M) | $|Zn^{2+}$(0.1 M) | $Zn_{(s)}$ will be _______. (Given $E^{o}_{cell} = -1.562$ V)

[WB JEEM 2015]

(A) – 1.532 V (B) – 1.503 V

(C) 1.532 V (D) – 3.06 V

7. Find the emf of the cell in which the following reaction takes place at 298 K.

$$Ni(s) + 2Ag^{+}\ (0.001\ M) \rightarrow Ni^{2+}\ (0.001\ M) + 2Ag(s)$$

(Given that $E^{o}_{cell} = 10.5$ V, $\frac{2.303RT}{F} = 0.059$ at 298 K) **[NEET (UG) 2022]**

(A) 0.9615 V (B) 1.05 V

(C) 1.0385 V (D) 1.385 V

8. The e.m.f. of the cell $Ag \mid Ag^{+}(0.1M) \parallel Ag^{+}(1M) \mid Ag$ at 298 K is _______.

(A) 0.0059 V (B) 0.059 V

(C) 5.9 V (D) 0.59 V

9. The standard e.m.f. for the given cell reaction, $Zn + Cu^{2+} \longrightarrow Cu + Zn^{2+}$ is 1.10 V at 25 °C. The e.m.f. for the cell reaction, when 0.1 M Cu^{2+} and 0.1 M Zn^{2+} solutions are used, at 25 °C is _______.

(A) 1.10 V (B) 0.110 V

(C) –1.10 V (D) –0.110 V

10. For the redox reaction,

$Zn(s) + Cu^{2+}\ (0.1M) \longrightarrow Zn^{2+}\ (1M) + Cu(s)$ taking place in a cell, E^{o}_{cell} is 1.10 volt. At 298 K, E_{cell} for the cell will be _______.

(A) 2.14 volt (B) 1.80 volt

(C) 1.07 volt (D) 1.13 volt

11. To find the standard potential of $M^{3+}|M$ electrode, the following cell is constituted:

$Pt \mid M \mid M^{3+}(0.001\ M) \parallel Ag^{+}(0.01\ M) \mid Ag$.

The emf of the cell is found to be 0.421 V at 298 K. The standard potential of half reaction $M^{3+} + 3e^{-} \rightarrow M$ at 298 K will be________.

(Given, $E^{o}_{Ag^{+}|Ag}$ at 298 K = 0.80 V)

(A) 0.38 V (B) 0.32 V

(C) 1.28 V (D) 0.66 V

12. If Zn^{2+} and Cu^{2+} concentrations are 0.1 M and 10^{-9} M respectively at 25 °C, the potential of the cell containing Zn/Zn^{2+} and Cu/Cu^{2+} electrodes is _______.

(Given $E^{o}_{Zn^{2+}/Zn} = -0.76$; $E^{o}_{Cu^{2+}/Cu} = +0.34$ V)

(A) 0.864 V (B) – 0.864 V

(C) 1.33 V (D) 1.10 V

13. The e.m.f. of the cell in the reaction

$$Zn_{(s)} + Ni^{2+}(C_1 = 1.0M) \rightleftharpoons Zn^{2+}(C_2 = 10M) + Ni_{(s)}$$

is found to be 0.5105 V at 298 K. The standard e.m.f. of the cell is _______.

(A) 0.5400 V (B) 0.4810 V

(C) 0.5696 V (D) – 0.5105 V

14. Under which of the following conditions E value of the cell for the cell reaction given is maximum?

$$\underset{}{Zn_{(S)}} + \underset{C_1}{Cu^{2+}_{(aq)}} \rightleftharpoons Cu + \underset{C_2}{Zn^{2+}_{(aq)}}$$

$\left(\frac{2.303\ RT}{F}\text{at } 298\ K = 0.059\ V,\ E^{o}_{Zn^{2+}|Zn} = -0.76\ V,\ E^{o}_{Cu^{2+}|Cu} = +0.37\ V\right)$

[AP EAMCET (Engg.) 2019]

(A) $C_1 = 0.1$ M, $C_2 = 0.01$ M

(B) $C_1 = 0.01$ M, $C_2 = 0.1$ M

(C) $C_1 = 0.1$ M, $C_2 = 0.2$ M

(D) $C_1 = 0.2$ M, $C_2 = 0.1$ M

15. Calculate the electrode potential at 298 K for $Zn|Zn^{2+}$ electrode in which the activity of zinc ions is 0.001 M and $E^{o}_{(Zn/Zn^{2+})}$ is –0.74 volts.

(A) 0. 38 volts (B) – 0.83 volts

(C) 0.40 volts (D) – 0.45 volts

16. Consider the following electrodes

$P = Zn^{2+}$ (0.0001 M) | Zn

$Q = Zn^{2+}$ (0.1 M) | Zn

$R = Zn^{2+}$ (0.01 M) | Zn

$S = Zn^{2+}$ (0.001 M) | Zn

$E^{o}_{Zn/Zn^{2+}} = -0.76$ V

Electrode potentials of the above electrodes in volts are in the order: **[KCET 2021]**

(A) P > S > R > Q

(B) S > R > Q > P

(C) Q > R > S > P

(D) P > Q > R > S

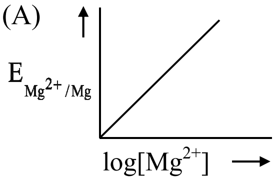
17. Electrode potential for Mg electrode varies according to the equation

$$E_{Mg^{2+}/Mg} = E^{o}_{Mg^{2+}/Mg} - \frac{0.059}{2}\ \log\frac{1}{[Mg^{2+}]}.$$

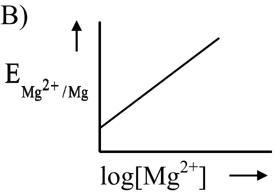
The graph of $E_{Mg^{2+}/Mg}$ versus log $[Mg^{2+}]$ is _______ .

[NCERT Exemplar]

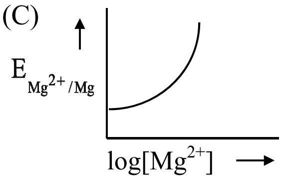
(A)



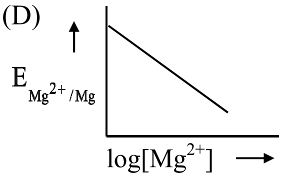
(B)



(C)



(D)



18. What will be the oxidation potential for the following hydrogen half-cell at 1 bar pressure and 25 °C temperature?

$\text{Pt}\,|\,H_{2(g,\ 1\ bar)}\,|\,HCl_{(aq,\ pH=3)}$ **[GUJ CET 2018]**

(A) 0.059 V (B) 0.188 V
(C) 0.177 V (D) 0.000 V

19. A hydrogen gas electrode is made by dipping platinum wire in a solution of HCl of pH = 10 and by passing hydrogen gas around the platinum wire at one atm pressure. The oxidation potential of electrode would be _______. **[NEET (UG) 2013]**

(A) 0.59 V (B) 0.118 V
(C) 1.18 V (D) 0.059 V

20. The hydrogen electrode is dipped in a solution of pH = 2 at 25 °C and 1 bar. The potential of the electrode would be _______.
(the value of 2.303 RT/F is 0.059 V)

(A) 0.118 V (B) – 0.118 V
(C) 0.087 V (D) 0.059 V

21. At pH = 1 and 1 bar pressure, potential of hydrogen electrode at 298 K is _______.

(A) – 0.59 V (B) 0.00 V
(C) – 0.059 V (D) – 0.295 V

22. If the pressure of hydrogen gas is increased from 1 atm to 100 atm, keeping the hydrogen ion concentration constant at 1 M, the voltage of hydrogen half-cell will be _______.

(A) –0.089 V (B) – 0.059 V
(C) –0.295 V (D) –0.118 V

23. The pressure of H_2 required to make the potential of H_2–electrode zero in pure water at 298 K is _______. **[NEET (UG) P-I 2016]**

(A) 10^{-10} atm (B) 10^{-4} atm
(C) 10^{-14} atm (D) 10^{-12} atm

24. The reduction potential of hydrogen half-cell at 298 K is negative if _______.

(A) $P(H_2) = 1$ atm and $[H^+] = 1$ M
(B) $P(H_2) = 1$ atm and $[H^+] = 2$ M
(C) $P(H_2) = 2$ atm and $[H^+] = 1$ M
(D) $P(H_2) = 2$ atm and $[H^+] = 2$ M

25. Consider the single electrode process $4H^+ + 4e^- \rightarrow 2H_2$ catalyzed by platinum black electrode in HCl electrolyte. The potential of the electrode is –0.059 versus standard hydrogen electrode. What is the concentration of the acid in the hydrogen half-cell if the H_2 pressure is 1 bar? **[TS-EAMCET 2017]**

(A) 1 M (B) 10 M
(C) 0.1 M (D) 0.01 M

26. What is the potential of a half-cell consisting of zinc electrode in 0.01 M $ZnSO_4$ solution at 25 °C ($E° = –0.763$ V)?

(A) –0.822 V (B) 8.22 V
(C) –0.528 V (D) 9.23 V

27. In the electrochemical cell
Zn | $ZnSO_4$(0.01 M) || $CuSO_4$(1.0 M) | Cu, the emf of this Daniel Cell is E_1. When the concentration of $ZnSO_4$ is changed to 1.0 M and that of $CuSO_4$ changed to 0.01 M, the emf changes to E_2. From the following, which one is the relationship between E_1 and E_2?

(Given, $\frac{RT}{F} = 0.059$) **[NEET (UG) 2017]**

(A) $E_1 < E_2$ (B) $E_1 > E_2$
(C) $E_2 = 0 \neq E_1$ (D) $E_1 = E_2$

28. In the cell reaction,

$Cu_{(s)} + 2Ag^+_{(aq)} \rightarrow Cu^{2+} + 2Ag_{(s)},\ E^{o}_{cell} = 0.46$ V.

By doubling the concentration of Cu^{2+}; E_{cell} _______.

(A) is doubled
(B) is halved
(C) increases but less than double
(D) decreases by a small fraction

29. In the galvanic cell
Zn | Zn^{2+} (C_1) || Cu^{2+} (C_2) | Cu
$E_{cell} - E^{o}_{cell} = +0.0591$ V.

The value of $\frac{C_1}{C_2}$ at 298 K is _______ .

[AP EAMCET (Med.) 2016]

(A) 10^{-1} (B) 10^{2} (C) 10^{-2} (D) 10

30. The e.m.f. of the cell
$Zn\,|\,Zn^{2+}(0.01M)\,||\,Fe^{2+}(0.001M)\,|\,Fe$ at 298 K is 0.2905 volts; then the value of equilibrium constant for the cell reaction is _______.

(A) $10^{\frac{0.32}{0.0295}}$ (B) $10^{\frac{0.32}{0.295}}$
(C) $10^{\frac{0.26}{0.0295}}$ (D) $10^{\frac{0.32}{0.0591}}$

31. For $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$; $E° = 1.33$ V. At $\left[Cr_2O_7^{-2}\right] = 4.5$ millimole, $[Cr^{+3}]$ = 15 millimole, E is 1.067 V. The pH of the solution is nearly equal to _______. **[KCET 2014]**

(A) 2 (B) 3 (C) 5 (D) 4

32. What will be the e.m.f. for the given cell $Pt\,|\,H_2(P_1)\,|\,H^+_{(aq)}\,||\,H_2(P_2)\,|\,Pt$?

(A) $\frac{RT}{F}\ln\frac{P_1}{P_2}$ (B) $\frac{RT}{2F}\ln\frac{P_1}{P_2}$

(C) $\frac{RT}{F}\ln\frac{P_2}{P_1}$ (D) $\frac{RT}{2F}\ln\frac{P_2}{P_1}$

33. If hydrogen electrode is dipped in 2 solutions of pH = 3 and pH = 6 and salt bridge is connected; the e.m.f. of resulting cell is _______.

(A) 0.177 V (B) 0.3 V

(C) 0.052 V (D) 0.104 V

34. What is the potential of a cell containing two hydrogen electrodes both having hydrogen gas at 1 bar pressure; the negative one in contact with 10^{-8} M H^+ and positive one in contact with 0.025 M H^+? [Given: log 0.025 = 1.602]

(A) 0.18 V (B) 0.28 V

(C) 0.38 V (D) 0.48 V

35. Which of the following is concentration cell? **[GUJ CET 2014]**

(A) $Cu_{(s)}\,|\,Cu^{2+}(aq, 1\text{ M})\,||\,Cu^{2+}(aq, 1\text{ M})\,|\,Cu_{(s)}$

(B) $Cu_{(s)}\,|\,Cu^{2+}(aq, 0.5\text{ M})\,||\,Cu^{2+}(aq, 0.5\text{ M})\,|\,Cu_{(s)}$

(C) $Zn_{(s)}\,|\,Zn^{2+}(aq, 0.5\text{ M})\,||\,Cu^{2+}(aq, 1\text{ M})\,|\,Cu_{(s)}$

(D) $^{\ominus}Pt\,|\,H_2(g, 1\text{ bar})\,|HCl(aq, 0.002\text{ M})\,||\,HCl(aq, 0.005\text{ M})|\,H_2(g, 1\text{ bar})\,|\,Pt^{\oplus}$

CONDUCTANCE IN ELECTROLYTIC SOLUTIONS

1. In the electronic conductor, the conduction of electricity occurs by the direct flow of ______.

(A) positive ions (B) negative ions

(C) electrons (D) all of these

2. At which temperature, ceramic materials behave as superconductors? **[GUJ CET 2019]**

(A) 150 K (B) 200 K

(C) 15 K (D) 0 K

3. Strong electrolytes are those which _______.

(A) dissolve readily in water

(B) conduct electricity

(C) dissociate into ions at high dilution

(D) completely dissociate into ions at all dilutions.

4. Which of the following conducts electricity?

(A) Fused NaCl (B) CO_2

(C) Br_2 (D) urea

5. The one that is a good conductor of electricity in solid state, among the following is _______.

(A) sodium chloride (B) diamond

(C) graphite (D) plastic rod

6. Aqueous solution of which of the following compounds is the best conductor of electric current? **[AIPMT Re-Test 2015]**

(A) Ammonia, NH_3

(B) Fructose, $C_6H_{12}O_6$

(C) Acetic acid, $C_2H_4O_2$

(D) Hydrochloric acid, HCl

7. Which of the following compounds will NOT undergo dissociation in aqueous solution?

(A) Sugar (B) Sodium Chloride

(C) Sodium Bromide (D) Sodium Acetate

8. Which of the following is a poor conductor of electricity?

(A) CH_3COONa (B) C_2H_5OH

(C) NaCl (D) KOH

9. Which one of the following is NOT a conductor of electricity?

(A) NaCl (aqueous) (B) NaCl (solid)

(C) NaCl (molten) (D) Ag metal

10. Which one of the following material conducts electricity?

(A) Crystalline sodium chloride

(B) Barium sulphate

(C) Fused potassium chloride

(D) Molten sulphur

11. The electrolytic conductance is a direct measure of _______.

(A) resistance (B) potential

(C) concentration (D) dissociation

12. Which of the following statement is INCORRECT with respect to metallic or electronic conductivity? **[GUJ CET 2014]**

(A) Metallic conductivity depends on the structure of metal and its characteristics.

(B) Metallic conductivity depends on the number of electrons in the valence shell of atom of metal.

(C) The electrical conductivity of metal increases with increase in temperature.

(D) There is no change in the structure of metal during electrical conduction.

13. The electrical conductance of a material is ________ proportional to its area of cross section and _______ proportional to its length.

(A) directly, directly

(B) directly, inversely

(C) inversely, directly

(D) inversely, inversely

14. 1 siemen = _______.

(A) 1 Ω
(B) $1\ \Omega^{-1}$
(C) 1 Ωm
(D) $1\ \Omega^{-1}m^{-1}$

15. The unit ohm^{-1} is used for _______.

(A) molar conductivity
(B) equivalent conductivity
(C) specific conductance
(D) conductance

16. Conductivity of an electrolytic solution and cell constant are related by _______.
[MHT CET 2019]

(A) $k = R \times \frac{l}{a}$
(B) $k = \frac{1}{R} \times \frac{1}{a}$
(C) $k = \frac{1}{R} \times \frac{a}{l}$
(D) $k = R \times \frac{a}{1}$

17. What is the SI unit of conductivity?
[MHT CET 2017]

(A) S m
(B) $S\ m^{-1}$
(C) $S\ m^2$
(D) $S\ m^{-2}$

18. Conductance (unit Siemens) is directly proportional to area of the vessel and is inversely proportional to the length of the vessel then the unit of the constant of proportionality is _______.

(A) $Sm\ mol^{-1}$
(B) Sm^{-1}
(C) $S^{-2}m^2\ mol$
(D) $S^2m^2\ mol^{-2}$

19. The specific conductance of four electrolytes in $ohm^{-1}\ cm^{-1}$ are given below. Which one of the following offers the highest resistance to passage of electric current ?

(A) 7.0×10^{-5}
(B) 9.2×10^{-9}
(C) 6.0×10^{-7}
(D) 4.0×10^{-8}

20. The unit of molar conductivity is _______.

(A) $\Omega^{-1}cm^{-2}mol^{-1}$
(B) $\Omega\ cm^{-2}mol^{-1}$
(C) $\Omega^{-1}cm^2mol^{-1}$
(D) $\Omega\ cm^2mol$

21. What is the molar conductivity of 0.20 M KCl solution if its conductivity is 0.0242 S cm^{-1} at 298 K?
[MHT CET 2021]

(A) 148.4 S $cm^2\ mol^{-1}$
(B) 82.6 S $cm^2\ mol^{-1}$
(C) 121.0 S $cm^2\ mol^{-1}$
(D) 484.0 S $cm^2\ mol^{-1}$

22. The unit of equivalent conductivity is _______.

(A) ohm cm
(B) $ohm^{-1}\ cm^2$(gm equivalent)$^{-1}$
(C) ohm cm^2 (gm equivalent)
(D) Scm^{-2}

23. Given $\frac{l}{a}$ = 0.5 cm^{-1}, R = 50 ohm, concentration = 1.0 N. The equivalent conductance of the electrolytic cell is _______.

(A) 10 $ohm^{-1}cm^2\ gm\,eq^{-1}$
(B) 20 $ohm^{-1}cm^2\ gm\,eq^{-1}$
(C) 300 $ohm^{-1}cm^2\ gm\,eq^{-1}$
(D) 100 $ohm^{-1}cm^2\ gm\,eq^{-1}$

24. At a particular temperature the ratio of equivalent conductance to specific conductance of a 0.01 N NaCl solution is _______.
[WB JEE 2015]

(A) $10^5\ cm^3\ eq^{-1}$
(B) $10^3\ cm^3\ eq^{-1}$
(C) $10\ cm^3\ eq^{-1}$
(D) $10^5\ cm^2\ eq^{-1}$

25. At a particular temperature, the ratio of molar conductance to specific conductance of 0.01M NaCl solution is _______. **[KCET 2018]**

(A) $10^5\ cm^3\ mol^{-1}$
(B) $10^3\ cm^3\ mol^{-1}$
(C) $10\ cm^3\ mol^{-1}$
(D) $10^5\ cm^2\ mol^{-1}$

MEASUREMENT OF CONDUCTIVITY

1. If '*l*' stands for the distance between the electrodes and 'a' stands for the area of cross section of the electrode, $\frac{l}{a}$ refers to _______.

(A) the degree of dissociation
(B) conductivity
(C) cell constant
(D) molar conductivity

2. The cell constant of a conductivity cell _______. **[NCERT Exemplar]**

(A) changes with change of electrolyte
(B) changes with change of concentration of electrolyte
(C) changes with temperature of electrolyte
(D) remains constant for a cell

3. The cell constant is the product of resistance and _______.

(A) conductance
(B) molar conductance
(C) specific conductance
(D) specific resistance

4. The unit of cell constant is _______.

(A) $ohm^{-1}\ cm^{-1}$
(B) ohm cm
(C) cm
(D) cm^{-1}

5. The two electrodes of platinum fitted in a conductance cell are 1.5 cm apart while the area of cross section of each electrode is 0.75 cm^2. The cell constant is _______.

(A) 0.2 cm^{-1}
(B) 0.5 cm^{-1}
(C) 0.125 cm^{-1}
(D) 2.0 cm^{-1}

6. What is the cell constant of $\frac{N}{10}$ KCl solution at 25 °C, if conductivity and resistance of a solution is 0.0112 Ω^{-1} cm^{-1} and 55.0 Ω respectively? **[MHT CET 2020]**

(A) 2.0 cm^{-1} (B) 0.491 cm^{-1}
(C) 0.2 cm^{-1} (D) 0.616 cm^{-1}

7. The resistance of 0.01 m KCl solution at 298 K is 1500 Ω. If the conductivity of 0.01 m KCl solution at 298 K is 0.146×10^{-3} S cm^{-1}. The cell constant of the conductivity cell in cm^{-1} is: **[KCET 2021]**

(A) 0.219 (B) 0.291
(C) 0.301 (D) 0.194

8. The conductivity of centimolar solution of KCl at 25°C is 0.0210 ohm^{-1} cm^{-1} and the resistance of the cell containing the solution at 25°C is 60 ohm. The value of cell constant is **[NEET (UG) 2023]**

(A) 3.28 cm^{-1} (B) 1.26 cm^{-1}
(C) 3.34 cm^{-1} (D) 1.34 cm^{-1}

9. The specific conductance of a 0.1 N KCl solution at 23 °C is 0.012 ohm^{-1} cm^{-1}. The resistance of cell containing the solution at the same temperature was found to be 55 ohm. The cell constant will be _______.

(A) 0.142 cm^{-1} (B) 0.66 cm^{-1}
(C) 0.918 cm^{-1} (D) 1.12 cm^{-1}

10. A conductivity cell has been calibrated with a 0.01 M 1:1 electrolyte solution (specific conductance, $\kappa = 1.25 \times 10^{-3}$ S cm^{-1}) in the cell and the measured resistance was 800 ohms at 25 °C. The cell constant will be _______ . **[WB JEE 2013]**

(A) 1.02 cm^{-1} (B) 0.102 cm^{-1}
(C) 1.00 cm^{-1} (D) 0.5 cm^{-1}

11. The distance between the electrodes of a conductivity cell is 0.98 cm and the cell constant is 0.5 cm^{-1}. Calculate the cross-sectional area of electrode. **[MHT CET 2019]**

(A) 1.96 cm^2 (B) 3.92 cm^2
(C) 0.4 cm^2 (D) 0.5 cm^2

12. Resistance of 0.2 M solution of an electrolyte is 50 Ω. The specific conductance of the solution is 1.4 S m^{-1}. The resistance of 0.5 M solution of the same electrolyte is 280 Ω. The molar conductivity of 0.5 M solution of the electrolyte in S m^2 mol^{-1} is _______. **[JEE (Main) 2014]**

(A) 5×10^{-4} (B) 5×10^{-3}
(C) 5×10^{3} (D) 5×10^{2}

13. The electrical resistance of a column of 0.05 M NaOH solution of diameter 1 cm and length 50 cm is 6.5×10^3 ohm. Its molar conductivity will be _______. **[AP EAPCET (Agri. & Pharm.) 2021]**

(A) 229.5 S cm^2 mol^{-1} (B) 196 S cm^2 mol^{-1}
(C) 149 S cm^2 mol^{-1} (D) 280 S cm^2 mol^{-1}

VARIATION OF CONDUCTIVITY AND MOLAR CONDUCTIVITY WITH CONCENTRATION

1. Which of the following increases with the increase in concentration of the solution?

(A) Conductance
(B) Specific conductance
(C) Equivalent conductance
(D) Molar conductance

2. Conductivity of a solution is directly proportional to _______.

(A) temperature of the solution
(B) number of ions present in the solution
(C) greater polarity of the solvent
(D) all of these

3. The molar conductivity is maximum for the solution of concentration _______.

(A) 0.001 M (B) 0.005 M
(C) 0.002 M (D) 0.004 M

4. The highest electrical conductivity of the following aqueous solutions is of _______.

(A) 0.1 M acetic acid
(B) 0.1 M chloroacetic acid
(C) 0.1 M fluoroacetic acid
(D) 0.1 M difluoroacetic acid

5. When a solution of an electrolyte is heated, the conductance of the solution, _______.

(A) increases because of the increased heat
(B) decreases because of the increased heat
(C) decreases because of the dissociation of the electrolyte is suppressed
(D) increases because the electrolyte is dissociated more

6. Which of the following statements about solutions of electrolytes is INCORRECT? **[NCERT Exemplar]**

(A) Conductivity of solution depends upon size of ions.
(B) Conductivity depends upon viscosity of solution.
(C) Conductivity does not depend upon solvation of ions present in solution.
(D) Conductivity of solution increases with temperature.

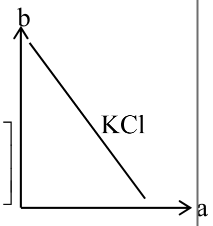
7. Molar conductivity of weak electrolytes increases with dilution. This is due to the _______.
(A) increase in the degree of ionization of the substance
(B) increase in the number of ions per unit volume
(C) increase in the molecular attractions
(D) increase in the degree of association

8. Molar conductance of an electrolyte increases with dilution according to the equation:
$\Lambda_m = \Lambda^{\circ}_{m} - A\sqrt{c}$
Which of the following statements are TRUE?
(i) This equation applies to both strong and weak electrolytes.
(ii) Value of the constant A depends upon the nature of the solvent.
(iii) Value of constant A is same for both $BaCl_2$ and $MgSO_4$.
(iv) Value of constant A is same for both $BaCl_2$ and $Mg(OH)_2$.
Choose the most appropriate answer from the options given below.
[NEET (UG) Manipur 2023]
(A) (ii) and (iii) only
(B) (ii) and (iv) only
(C) (i) and (ii) only
(D) (i), (ii) and (iii) only

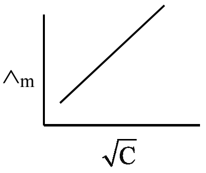
9. Which one of the following solutions will have highest conductivity? **[WB JEE 2021]**
(A) 0.1 M CH_3COOH (B) 0.1 M NaCl
(C) 0.1 M KNO_3 (D) 0.1 M HCl

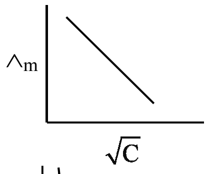
10. The equivalent conductance of NaCl at concentration C and at infinite dilution are λ_C and λ_∞ respectively. The CORRECT relationship between λ_C and λ_∞, is given as _______.
(where the constant B is positive)
[JEE (Main) 2014]
(A) $\lambda_C = \lambda_\infty + (B)C$
(B) $\lambda_C = \lambda_\infty - (B)C$
(C) $\lambda_C = \lambda_\infty - (B)\sqrt{C}$
(D) $\lambda_C = \lambda_\infty + (B)\sqrt{C}$

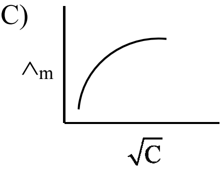
11. The CORRECT unit of the slope given in the following plot for the strong electrolyte is _______.
$\left[\begin{array}{l} a = (\text{concentration})^{\frac{1}{2}}, \\ b = \text{molar conductivity of solution} \end{array}\right]$

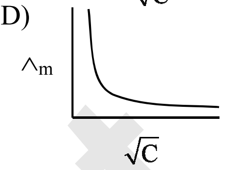


[TS EAMCET (Med.) 2019]
(A) $\dfrac{S\,cm^2\,mol}{(mol\,L^{-1})^{\frac{1}{2}}}$ (B) $\dfrac{S\,cm^2\,L^{-1}}{(mol)^{\frac{1}{2}}}$
(C) $\dfrac{S\,cm^2\,mol^{-1}}{(mol\,L^{-1})^{\frac{1}{2}}}$ (D) $\dfrac{S\,cm^3\,mol^{-2}}{(mol\,L^{-1})^{2}}$

12. Which one is CORRECT for acetic acid?
[Assam CEE 2017]
(A)


(B)


(C)


(D)


13. **Assertion:** $\wedge_m$ for weak electrolytes increases steeply on increasing dilution.
Reason: For weak electrolytes, α increases as dilution of solution is increased.
(A) Assertion and Reason are true. Reason is correct explanation of Assertion.
(B) Assertion and Reason are true. Reason is not the correct explanation of Assertion.
(C) Assertion is true, Reason is false.
(D) Assertion is false, Reason is true.

14. The pair of electrolytes that possess same value for the constant (A) in the Debye–Huckel–Onsagar equation, $\lambda_m = \lambda^{\circ}_{m} - A\sqrt{C}$ is _______. **[KCET 2020]**
(A) NH_4Cl, NaBr (B) NaBr, $MgSO_4$
(C) NaCl, $CaCl_2$ (D) $MgSO_4$, Na_2SO_4

KOHLRAUSCH'S LAW AND ITS APPLICATIONS

1. According to Kohlrausch's law for infinite dilution, the equivalent conductance of the electrolyte is equal to _______.
(A) sum of the equivalent conductance of the cations and anions present in it
(B) difference of the equivalent conductance of cations and anions present in it
(C) the ratio of the conductance of the cations to the anions present in it
(D) none of these

2. Molar conductivity of KCl at infinite dilution is 149.9 and that of NaCl is 126.5 $ohm^{-1}cm^2mol^{-1}$. Although at infinite dilution, each electrolyte is completely ionized, yet their molar conductivities differ from each other. This is due to the fact that _______.
(A) degree of ionization of KCl is greater than the degree of ionization of NaCl
(B) molar conductivity of K^+ ion at infinite dilution is more than that of Na^+ ion
(C) degree of ionization of NaCl is greater than the degree of ionization of KCl
(D) molar conductivity of Na^+ ion at infinite dilution is more than of K^+ ion

3. $\wedge^{o}_{m(NH_4OH)}$ is the equal to _______.

[NCERT Exemplar]

(A) $\wedge^{o}_{m(NH_4OH)} + \wedge^{o}_{m(NH_4Cl)} - \wedge^{o}_{(HCl)}$

(B) $\wedge^{o}_{m(NH_4Cl)} + \wedge^{o}_{m(NaOH)} - \wedge^{o}_{(NaCl)}$

(C) $\wedge^{o}_{m(NH_4Cl)} + \wedge^{o}_{m(NaCl)} - \wedge^{o}_{(NaOH)}$

(D) $\wedge^{o}_{m(NaOH)} + \wedge^{o}_{m(NaCl)} - \wedge^{o}_{(NH_4Cl)}$

4. What is the molar conductivity at infinite dilution of $CaCl_2$, if the molar conductivity of Ca^{2+} ion and Cl^{-} ion at infinite dilution is 119 and 71 Ω^{-1} cm^2 mol^{-1}? **[MHT CET 2020]**

(A) 431.0 Ω^{-1} cm^2 mol^{-1}

(B) 341.0 Ω^{-1} cm^2 mol^{-1}

(C) 261.0 Ω^{-1} cm^2 mol^{-1}

(D) 126.0 Ω^{-1} cm^2 mol^{-1}

5. The limiting molar conductivities $\wedge^{o}_{m}$ for NaCl, KBr and KCl are 126, 152 and 150 S cm^{2}mol^{-1}, respectively. The $\wedge^{o}_{m}$ for NaBr is _______.

(A) 278 S cm^{2}mol^{-1} (B) 176 S cm^{2}mol^{-1}

(C) 128 S cm^{2}mol^{-1} (D) 302 S cm^{2}mol^{-1}

6. The molar conductance of NaCl, HCl and CH_3COONa at infinite dilution are 126.45, 426.16 and 91.0 S cm^2 mol^{-1} respectively. The molar conductance of CH_3COOH at infinite dilution is _______. Choose the right option for answer. **[NEET (UG) 2021]**

(A) 390.71 S cm^2 mol^{-1}

(B) 698.28 S cm^2 mol^{-1}

(C) 540.48 S cm^2 mol^{-1}

(D) 201.28 S cm^2 mol^{-1}

7. Following limiting molar conductivities are given as

$\wedge^{o}_{m}$ (H_2SO_4) = x S cm^2 mol^{-1}

$\wedge^{o}_{m}$ (K_2SO_4) = y S cm^2 mol^{-1}

$\wedge^{o}_{m}$ (CH_3COOK) = z S cm^2 mol^{-1}

$\wedge^{o}_{m}$ (in S cm^2 mol^{-1}) for CH_3COOH will be _______. **[NEET (UG) Odisha 2019]**

(A) $\left(\frac{x-y}{2}\right) + z$ (B) $x - y + 2z$

(C) $x + y + z$ (D) $x - y + z$

8. The equivalent conductance of two strong electrolytes at infinite dilution in H_2O (where ions move freely through a solution) at 25 °C are $\wedge^{o}_{e(CH_3COONa)}$ = 91.0 S cm^2 equiv^{-1} and $\wedge^{o}_{e(HCl)}$ = 426.2 S cm^2 equiv^{-1} What additional information /quantity is needed to calculate $\wedge^{o}$ of an aqueous solution of acetic acid.

(A) $\wedge^{0}_{e}$ of chloroacetic acid

(B) $\wedge^{0}_{e}$ of NaCl

(C) $\wedge^{0}_{e}$ of CH_3COOK

(D) $\lambda^{o}_{e\,(H^+)}$

9. Equivalent conductivity at infinite dilution for sodium-potassium oxalate $((COO^{-})_2Na^{+}K^{+})$ will be _______. (given, molar conductivities of oxalate, K^{+} and Na^{+} ions at infinite dilution are 148.2, 73.5, 50.1 S cm^2 mol^{-1}, respectively) **[WB JEE 2013]**

(A) 271.8 S cm^2 eq^{-1}

(B) 67.95 S cm^2 eq^{-1}

(C) 543.6 S cm^2 eq^{-1}

(D) 135.9 S cm^2 eq^{-1}

10. The CORRECT order of equivalent conductance at infinite dilution of LiCl, NaCl and KCl is _______. **[BCECE (Stage 2) 2016]**

(A) LiCl > NaCl > KCl

(B) KCl > NaCl > LiCl

(C) NaCl > KCl > LiCl

(D) LiCl > KCl > NaCl

11. The molar conductance of 0.001 M acetic acid is 50 ohm^{-1} cm^2 mol^{-1}. The limiting molar conductance is 250 ohm^{-1} cm^2 mol^{-1}. What is its degree of ionization?

(A) 0.2 (B) 2 (C) 20 (D) 0.5

12. If equivalent conductance of 1M benzoic acid is 12.8 ohm^{-1} cm^2 eq^{-1} and that of benzoate ion and H^{+} ion at infinite dilution are 42 and 288.42 ohm^{-1} cm^2 eq^{-1} respectively, its degree of dissociation is _______.

(A) 38 (B) 39

(C) 0.35 (D) 0.038

13. At 25 °C molar conductance of 0.1 molar aqueous solution of ammonium hydroxide is 9.54 ohm^{-1} cm^2 mol^{-1} and at infinite dilution its molar conductance is 238 ohm^{-1} cm^2 mol^{-1}. The degree or ionization of ammonium hydroxide at same concentration and temperature is

(A) 20.800 % (B) 4.008 %

(C) 40.800 % (D) 2.080 %

14. The molar conductivity of 0.007 M acetic acid is 20 S cm^2 mol^{-1}. What is the dissociation constant of acetic acid? Choose the CORRECT option.

$$\lambda^{o}_{H^+} = 350\ \text{S cm}^2\ \text{mol}^{-1},\ \lambda^{o}_{CH_3COO^-} = 50\ \text{S cm}^2\ \text{mol}^{-1}$$

[NEET (UG) 2021]

(A) 2.50×10^{-4} mol L^{-1}

(B) 1.75×10^{-5} mol L^{-1}

(C) 2.50×10^{-5} mol L^{-1}

(D) 1.75×10^{-4} mol L^{-1}

15. The molar conductivity of 0.01 M solution of acetic acid is 16.6 Ω^{-1} cm^2 mol^{-1} and at infinite dilution, it is 390.7 Ω^{-1} cm^2 mol^{-1}. If the degree of dissociation of acetic acid in 0.01 M solution is 0.0425, what is its dissociation constant?

(A) 1.50×10^{-5} (B) 1.89×10^{-9}
(C) 1.89×10^{-5} (D) 2.89×10^{-7}

ELECTROLYTIC CELLS AND ELECTROLYSIS

1. Electrolytic cell is used to convert _______.
(A) chemical energy to electrical energy
(B) electrical energy to chemical energy
(C) chemical energy to mechanical energy
(D) electrical energy to mechanical energy

2. In the electrolytic cell, flow of electrons is from _______.
(A) cathode to anode in solution
(B) anode to cathode through external supply
(C) cathode to anode through external supply
(D) anode to cathode through internal supply

3. If the current is passed into the solution of an electrolyte, _______.
(A) anions move towards anode, cations towards cathode
(B) anions and cations both move towards anode
(C) anions move towards cathode, cations towards anode
(D) no movement of ions takes place

4. During electrolysis, electrons are _______.
(A) lost
(B) gained
(C) gained by cations and lost by anions
(D) gained by anions and lost by cations

5. Electrolysis involves oxidation and reduction at _______.
(A) anode and cathode respectively
(B) cathode and anode respectively
(C) at both the electrodes
(D) None of these

6. During electrolysis, the species discharged at cathode are _______.
(A) ions (B) cation
(C) anion (D) all of these

7. During the electrolysis of fused NaCl, which of the following reactions occurs at anode?
(A) Chloride ions are oxidized.
(B) Chloride ions are reduced.
(C) Sodium ions are oxidized.
(D) Sodium ions are reduced.

8. In the electrolysis of aqueous sodium chloride solution, which of the half-cell reaction will occur at anode?
[NCERT Exemplar; KCET 2017]

(A) $Na^+_{(aq)} + e^- \longrightarrow Na_{(s)}$; $E^o_{cell} = -2.71V$

(B) $2H_2O_{(l)} \longrightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$; $E^o_{cell} = 1.23V$

(C) $H^+_{(aq)} + e^- \longrightarrow \frac{1}{2}H_{2(g)}$; $E^o_{cell} = 0.00$ V

(D) $Cl^-_{(aq)} \longrightarrow \frac{1}{2}Cl_{2(g)} + e^-$; $E^o_{cell} = 1.36$ V

9. The resulting solution obtained at the end of electrolysis of concentrated aqueous solution of NaCl _______. **[GUJ CET 2015]**
(A) turns red litmus into blue
(B) turns blue litmus into red
(C) remains colourless with phenolphthalein
(D) the colour of red or blue litmus does not change

10. What will happen during the electrolysis of aqueous solution of $CuSO_4$ by using platinum electrodes?
(A) Copper will deposit at cathode.
(B) Copper will deposit at anode.
(C) Hydrogen will be released at anode.
(D) Copper will dissolve at anode.

11. The platinum electrodes were immersed in a solution of cupric sulphate and electric current passed through the solution. After some time it was found that colour from copper sulphate disappeared with evolution of gas at the electrode. The colourless solution contains _______.
(A) platinum sulphate
(B) copper hydroxide
(C) copper sulphate
(D) sulphuric acid

12. An aqueous solution of $CuSO_4$ is subjected to electrolysis using inert electrodes. The pH of the solution will _______. **[KCET 2019]**
(A) decrease
(B) increase or decrease depending on the strength of the current
(C) increase
(D) remains unchanged

13. If electrolysis of aqueous $CuSO_4$ solution is carried out using Cu-electrodes, the reaction taking place at the anode is _______.
[WB JEE 2019]

(A) $H^+ + e^- \longrightarrow H$

(B) $Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$

(C) $SO^{2-}_{4(aq)} - 2e^- \longrightarrow SO_4$

(D) $Cu_{(s)} - 2e^- \longrightarrow Cu^{2+}_{(aq)}$

14. E^{o}_{cell} for some half-cell reactions are given below. On the basis of these mark the CORRECT answer.

i. $H^{+}(aq) + e^{-} \longrightarrow \frac{1}{2}H_2(g)$; $E^{o}_{cell} = 0.00$ V

ii. $2H_2O(l) \longrightarrow O_2(g) + 4H^{+}(aq) + 4e^{-}$; $E^{o}_{cell} = 1.23$ V

iii. $2SO_4^{2-}(aq) \longrightarrow S_2O_8^{2-}(aq) + 2e^{-}$; $E^{o}_{cell} = 1.96$ V

(A) In dilute sulphuric acid solution, hydrogen will be liberated at cathode.
(B) In concentrated sulphuric acid solution, water will be oxidized at anode.
(C) In dilute sulphuric acid solution, water will be reduced at cathode.
(D) In dilute sulphuric acid solution, SO_4^{2-} ion will be oxidized to tetrathionate ion at anode.

15. On electrolysis of dil. sulphuric acid using Platinum (Pt) electrode, the product obtained at anode will be _______. **[NEET (UG) P-I 2020]**

(A) Oxygen gas (B) H_2S gas
(C) SO_2 gas (D) Hydrogen gas

16. Which of the following statements is CORRECT for electrolysis of dilute H_2SO_4? **[BCECE (Stage 2) 2016]**

(A) O_2 is liberated on cathode.
(B) H_2 is liberated on anode.
(C) O_2 is liberated on anode.
(D) SO_2 is liberated on cathode.

17. The products obtained at the cathode and anode respectively during the electrolysis of aqueous K_2SO_4 solution using platinum electrodes are _______. **[AP EAMCET (Engg.) 2016]**

(A) O_2, H_2 (B) H_2, O_2
(C) H_2, SO_2 (D) K, SO_2

18. Which of the following compound does not give oxygen at anode and hydrogen at cathode during the electrolysis of their dilute aqueous solution?

(A) Na_2SO_4 (B) $AgNO_3$
(C) H_2SO_4 (D) $ZnSO_4$

19. Some statements are given below:

i. The electrolytic conduction is due to the migration of ions through the electrolyte.
ii. The passage of electric current through the electrolyte leads to the chemical changes.
iii. Electrolysis is a spontaneous redox reaction.
iv. Sodium chloride conducts electricity only in its aqueous solutions.

Among the above, the CORRECT statements are _______.

(A) (i) and (ii) (B) (i), (ii) and (iii)
(C) (i), (ii) and (iv) (D) all of these

FARADAY'S LAWS OF ELECTROLYSIS

1. The unit of electrochemical equivalent is _______.

(A) gram (B) gram/ampere
(C) gram/coulomb (D) coulomb/gram

2. During the electrolysis of an electrolyte, the number of ions produced, is directly proportional to the _______.

(A) time consumed
(B) electrochemical equivalent of electrolysis
(C) quantity of electricity passed
(D) mass of electrons

3. One Faraday = 96,500 _____.

(A) C mol (B) C mol^{-1}
(C) mol C^{-1} (D) mol^{-1} C^{-1}

4. The number of electrons delivered at the cathode during electrolysis by a current of 1 ampere in 60 seconds is _______. (charge on electron = 1.60×10^{-19} C) **[NEET (UG) P-II 2016]**

(A) 7.48×10^{23} (B) 6×10^{23}
(C) 6×10^{20} (D) 3.75×10^{20}

5. The number of moles of electrons passed when current of 2 A is passed through an solution of electrolyte for 20 minutes is _______. **[MHT CET 2018]**

(A) 4.1×10^{-4} mol e^{-}
(B) 1.24×10^{-2} mol e^{-}
(C) 2.487×10^{-2} mol e^{-}
(D) 2.487×10^{-1} mol e^{-}

6. The charge carried by 1 millimole of M^{n+} ions is 193 coulombs. The value of n is _______. **[WB JEE 2019]**

(A) 1 (B) 2 (C) 3 (D) 4

7. Aluminium oxide may be electrolysed at 1000°C to furnish aluminium metal (At. Mass = 27 u; 1 Faraday = 96,500 Coulombs). The cathode reaction is $Al^{3+} + 3e^{-} \rightarrow Al$
To prepare 5.12 kg of aluminium metal by this method, would require _______.

(A) 5.49×10^{7} C of electricity
(B) 1.83×10^{7} C of electricity
(C) 5.49×10^{4} C of electricity
(D) 5.49×10^{1} C of electricity

8. The number of coulombs required to reduce 12.3 g of nitrobenzene to aniline is _______.

(A) 115800 C (B) 5790 C
(C) 28950 C (D) 57900 C

9. The required charge for one equivalent weight of silver deposited on cathode is _______.

(A) $9.65\times10^{7}C$ (B) $9.65\times10^{4}C$
(C) $9.65\times10^{3}C$ (D) $9.65\times10^{5}C$

10. The quantity of electricity needed to liberate 0.5 gram equivalent of an element is _______.
(A) 48250 Faradays
(B) 48250 Coulombs
(C) 193000 Faradays
(D) 193000 Coulombs

11. How many coulombs of electricity are required for the oxidation of one mol of water to dioxygen? **[KCET 2015]**
(A) 9.65×10^4 C (B) 1.93×10^4 C
(C) 1.93×10^5 C (D) 19.3×10^5 C

12. How many coulombs of electricity is required to deposit 0.5 g of calcium metal (Molar mass = 40.0 g mol^{-1}) from calcium ions? **[MHT CET 2021]**
(A) 2412.5 C (B) 3612.5 C
(C) 2214.0 C (D) 3302.0 C

13. The number of Faradays (F) required to produced 20 g of calcium from molten $CaCl_2$ (Atomic mass of Ca = 40 g mol^{-1}) is _______. **[NEET (UG) P-I 2020]**
(A) 2 (B) 3 (C) 4 (D) 1

14. What amount of electricity can deposit 1 mole of Al metal at cathode when passed through molten $AlCl_3$? **[WB JEE 2018]**
(A) 0.3 F (B) 1 F (C) 3 F (D) 0.5 F

15. Amount of electricity that can deposit 108 g of silver from $AgNO_3$ solution is _______.
(A) 1 ampere (B) 1 coulomb
(C) 1 Faraday (D) 0.1 Faraday

16. The amount of current in Faraday required for the reduction of 1 mol of $Cr_2O_7^{2-}$ ions to Cr^{3+} is _______. **[KCET 2016; WB JEE 2017]**
(A) 1 F (B) 2 F (C) 6 F (D) 4 F

17. The charge required for the reduction of 1 mol of MnO_4^- to MnO_2 is _______. **[KCET 2018]**
(A) 1 F (B) 3 F (C) 5 F (D) 7 F

18. To deposit 0.6354 g of copper by electrolysis of aqueous cupric sulphate solution, the amount of electricity required (in coulombs) is _______.
(A) 9650 (B) 4825
(C) 3860 (D) 1930

19. When an electric current is passed through acidified water, 112 mL of hydrogen gas at S.T.P. was collected at the cathode in 965 seconds. The current passed in ampere is _______.
(A) 1.0 (B) 0.5 (C) 0.1 (D) 2.0

20. When 0.1 mol MnO_4^{2-} is oxidized, the quantity of electricity required to completely oxidise MnO_4^{2-} to MnO_4^- is ________. **[AIPMT 2014]**
(A) 96500 C (B) 2×96500 C
(C) 9650 C (D) 96.50 C

21. The current in a given wire is 1.8 A. The number of coulombs that flow in 1.36 minutes will be _______.
(A) 100 C (B) 147 C
(C) 247 C (D) 347 C

22. The number of coulombs required to liberate 0.224 dm^3 of chlorine at 0 °C and 1 atm pressure is _______.
(A) 2×965 (B) $\frac{965}{2}$
(C) 965 (D) 9,650

23. A current of 19296 C is passed through an aqueous solution of copper sulphate using copper electrodes. What is the mass (in g) of copper deposited at the cathode? (Molar mass of Cu = 63.5 g mol^{-1}) **[TS EAMCET (Engg.) 2019]**
(A) 3.17 (B) 1.58
(C) 6.35 (D) 0.79

24. Two faradays of electricity is passed through a solution of $CuSO_4$. The mass of copper deposited at the cathode is _______. (At. mass of Cu = 63.5 amu) **[JEE (Main) 2015]**
(A) 0 g (B) 63.5 g
(C) 2 g (D) 127 g

25. Three faradays electricity was passed through an aqueous solution of iron(II) bromide. The weight of iron metal (at. wt. = 56 amu) deposited at the cathode (in gm) is _______.
(A) 56 (B) 84 (C) 112 (D) 168

26. On passing 0.1 Faraday of electricity through aluminium chloride, the amount of aluminium metal deposited on cathode is _______.
(Al = 27 amu)
(A) 0.9 g (B) 0.3 g
(C) 0.27 g (D) 2.7 g

27. When 0.04 faraday of electricity is passed through molten $CaSO_4$, then the weight of Ca metal deposited at the cathode is _______.
(A) 0.2 g (B) 0.4 g (C) 0.6 g (D) 0.8 g

28. Solutions of $CuSO_4$ and $AgNO_3$ were electrolyzed with a current of 1.93 amperes for 500 seconds separately. The amount of copper and silver deposited at cathode respectively in g are _______. **[AP EAMCET (Med.) 2019]**
(A) 0.63, 0.54 (B) 0.315, 0.54
(C) 0.315, 1.08 (D) 1.08, 0.315

29. When a current of 10 A is passes through molten $AlCl_3$ for 1.608 minutes. The mass of Al deposited will be _______. [Atomic mass of Al = 27 g mol^{-1}] **[AP EAPCET (Engg.) 2021]**
(A) 0.09 g (B) 0.81 g
(C) 1.35 g (D) 0.27 g

30. By passing electric current, $NaClO_3$ is converted into $NaClO_4$ according to the following equation:

$NaClO_3 + H_2O \longrightarrow NaClO_4 + H_2$

How many moles of $NaClO_4$ will be formed when three faradays of charge is passed through $NaClO_3$? **[KCET 2017]**

(A) 0.75 mol (B) 1.0 mol
(C) 1.5 mol (D) 3.0 mol

31. One litre solution of $MgCl_2$ is electrolyzed completely by passing a current of 1 A for 16 min 5 sec. The original concentration of $MgCl_2$ solution was _______.
(Atomic mass of Mg = 24) **[KCET 2019]**

(A) 0.5×10^{-3} M (B) 1.0×10^{-2} M
(C) 5×10^{-3} M (D) 5×10^{-2} M

32. On electrolysis of aqueous solution of a halide of a metal 'M' by passing 1.5 ampere current for 10 minutes deposits 0.2938 g of metal. If the atomic mass of the metal is 63 g/mol, then what will be the formula of the metal halide? **[GUJ CET 2020]**

(A) MCl (B) MCl_3
(C) MCl_2 (D) MCl_4

33. A current of 2.0 A is passed for 5 hours through a molten metal salt and it deposits 22.2 g of metal (At. wt. =177). The oxidation state of the metal in the metal salt is _______.

(A) + 1 (B) + 2 (C) + 3 (D) + 4

34. What is the amount of chlorine evolved when 2 amperes of current is passed for 30 minutes in an aqueous solution of NaCl?

(A) 66 g (B) 1.32 g
(C) 33 g (D) 99 g

35. During the electrolysis of molten sodium chloride, the time required to produce 0.10 mol of chlorine gas using a current of 3 amperes is _______ . **[NEET (UG) P-II 2016]**

(A) 330 minutes (B) 55 minutes
(C) 110 minutes (D) 220 minutes

36. In the electrolysis of water, one Faraday of electrical energy would evolve _______.

(A) one mole of oxygen
(B) one g atom of oxygen
(C) 8 g of oxygen
(D) 22.4 litres of oxygen

37. When the same quantity of electricity is passed through the solution of different electrolytes in series, the amounts of product obtained are proportional to their _______.

(A) empirical mass
(B) chemical equivalents
(C) moles
(D) formula mass/96500

38. On passing one Faraday of electricity through the electrolytic cells containing Ag^+, Ni^{+2} and Cr^{+3} ions solution, the deposited amounts of Ag (At. wt. = 108), Ni (At. wt. = 59) and Cr (At. wt. = 52) are ______.

	Ag	Ni	Cr
(A)	108 g	29.5 g	17.3 g
(B)	108 g	59.0 g	52.0 g
(C)	108.0 g	108.0 g	108.0 g
(D)	108 g	117.5 g	166.0 g

39. The quantity of electricity needed to separately electrolyze 1 M solution of $ZnSO_4$, $AlCl_3$ and $AgNO_3$ completely is in the ratio of _______. **[WB JEEM 2014]**

(A) 2 : 3 : 1 (B) 2 : 1 : 1
(C) 2 : 1 : 3 (D) 2 : 2 : 1

40. A certain current liberated 0.504 g of hydrogen in 2 hours. How many grams of copper can be liberated by the same current flowing for the same time in a copper sulphate solution? (Atomic weight of Cu = 63.5 amu)

(A) 12.7 g (B) 16.0 g
(C) 31.8 g (D) 63.5 g

41. Two electrolytic cells, one containing acidified ferrous chloride and another acidified ferric chloride are connected in series. The ratio of iron deposited at cathodes in the two cells when electricity is passed through the cells will be _______.

(A) 3 : 1 (B) 2 : 1
(C) 1 : 1 (D) 3 : 2

42. Two electrolytic cells containing molten solutions of nickel chloride and aluminium chloride are connected in series. If same amount of electric current is passed through them, what will be the weight of nickel obtained when 18 g of aluminium is obtained?
(Al = 27 g mol^{-1}, Ni = 58.5 g mol^{-1}) **[GUJ CET 2015]**

(A) 58.5 g (B) 117 g
(C) 29.25 g (D) 5.85 g

43. The weight of silver (atomic weight = 108) displaced by a quantity of electricity which displaces 5600 mL of O_2 at STP will be ________. **[AIPMT 2014]**

(A) 5.4 g (B) 10.8 g
(C) 54.0 g (D) 108.0 g

44. In the electrolysis of acidulated water, a current of 0.5 A is passed through the cell for 30 minutes. The ratio of weight of H_2 and O_2 evolved is _______.

(A) 0.125 (B) 1.25
(C) 7.99 (D) 0.5

45. 4.5 g of aluminium (at mass 27 amu) is deposited at cathode from Al^{3+} solution by a certain quantity of electric charge. The volume of hydrogen produced at STP from H^+ ions in solution by the same quantity of electric charge will be _______.

(A) 22.4 L (B) 44.8 L
(C) 5.6 L (D) 11.2 L

46. Impure copper containing Fe, Au, Ag as impurities is electrolytically refined. A current of 140 A for 482.5 s decreased the mass of the anode by 22.26 g and increased the mass of cathode by 22.011 g. Percentage of iron in impure copper is _______.
(Given molar mass Fe = 55.5 g mol^{-1}, molar mass Cu = 63.54 g mol^{-1}) **[KCET 2014]**

(A) 0.95 (B) 0.85
(C) 0.97 (D) 0.90

TYPES OF CELLS OR BATTERIES

1. The type of a battery in which the reaction occurs only once and then battery becomes dead over a period of time is known as _____.

(A) primary cell
(B) secondary cell
(C) lead storage battery
(D) nickel cadmium cell

2. In dry cell, what acts as negative electrode? **[MHT CET 2016]**

(A) Zinc
(B) Graphite
(C) Ammonium chloride
(D) Manganese dioxide

3. In dry cell, the reaction which takes place at the zinc anode is _______.

(A) $Zn^{2+} + 2e^- \rightarrow Zn(s)$
(B) $Zn(s) \rightarrow Zn^{2+} + 2e^-$
(C) $Mn^{2+} + 2e^- \rightarrow Mn(s)$
(D) $Mn(s) \rightarrow Mn^+ + e^-$

4. **Assertion**: A rod of graphite is used as the cathode in ordinary dry cell batteries.
Reason: Graphite is a good conductor of electricity.
Select the CORRECT answer from the following.

(A) Assertion and Reason are true. Reason is correct explanation of Assertion.
(B) Assertion and Reason are true. Reason is not the correct explanation of Assertion.
(C) Assertion is true, Reason is false.
(D) Assertion is false, Reason is true.

5. Dry cells are used in _______.

(A) tape recorders (B) torches
(C) flash lights (D) all of these.

6. Which colourless gas evolves, when NH_4Cl reacts with zinc in a dry cell battery?

(A) HCl (B) N_2 (C) H_2 (D) Cl_2

7. **Assertion:** The cell potential of a mercury cell does not remain constant during its life time.
Reason: The overall cell reaction does not involve any ion in solution.

(A) Assertion and Reason are true. Reason is correct explanation of Assertion.
(B) Assertion and Reason are true. Reason is not the correct explanation of Assertion.
(C) Assertion is true, Reason is false.
(D) Assertion is false, Reason is true.

8. A secondary cell is one that _______. **[KCET 2016]**

(A) can be recharged
(B) can be recharged by passing current through it in the same direction
(C) can be recharged by passing current through it in the opposite direction
(D) cannot be recharged

9. The acid used in lead storage battery is _______.

(A) H_2SO_4 (B) H_3PO_4
(C) HCl (D) HNO_3

10. What is the density of solution of sulphuric acid used an electrolyte in lead accumulator? **[MHT CET 2018]**

(A) 1.5 g mL^{-1} (B) 1.2 g mL^{-1}
(C) 1.8 g mL^{-1} (D) 2.0 g mL^{-1}

11. When a lead storage battery is discharged, _______.

(A) SO_2 is evolved
(B) lead sulphate is consumed
(C) lead sulphite is consumed
(D) sulphuric acid is consumed

12. Which among the following equations represents the reduction reaction taking place in lead accumulator at positive electrode, while it is being used as a source of electrical energy? **[MHT CET 2017]**

(A) $Pb \rightarrow Pb^{2+}$ (B) $Pb^{4+} \rightarrow Pb$
(C) $Pb^{2+} \rightarrow Pb$ (D) $Pb^{4+} \rightarrow Pb^{2+}$

13. A lead storage battery is discharged. During the charging of this battery, the reaction that occurs at anode is _______. **[TS EAMCET (Engg.) 2015]**

(A) $PbSO_{4(s)} + 2e^- \longrightarrow Pb_{(s)} + SO^{2-}_{4\ (aq)}$
(B) $PbSO_{4(s)} + 2H_2O_{(l)} \longrightarrow PbO_{2(s)} + SO^{2-}_{4(aq)} + 4H^+_{(aq)} + 2e^-$
(C) $PbSO_{4(s)} \longrightarrow Pb^{2+}_{(aq)} + SO^{2-}_{4(aq)}$
(D) $PbSO_{4(s)} + 2H_2O_{(l)} + 2e^- \longrightarrow PbO_{2(s)} + SO^{2-}_{4(aq)} + 4H^+_{(aq)}$

14. While charging the lead storage battery, ________. **[NCERT Exemplar; KCET 2015]**
(A) $PbSO_4$ on anode is reduced to Pb
(B) $PbSO_4$ on cathode is reduced to Pb
(C) $PbSO_4$ on cathode is oxidized to Pb
(D) $PbSO_4$ on anode is oxidized to Pb

15. When lead accumulator is getting charged, it is _______.
(A) an electrolytic cell
(B) a galvanic cell
(C) a Daniel cell
(D) none of these

16. The ______ cell is used in electronic watches, calculators and photographic equipments.
(A) Daniel cell
(B) nickel cadmium cell
(C) Leclanche dry cell
(D) fuel

17. Nickel cadmium cell consists of _______ cathode.
(A) cadmium (B) NiO_2
(C) carbon (D) Pb

18. Nickel cadmium cells are _______.
(A) primary cells
(B) secondary cells
(C) fuel cell
(D) all of these

19. Nickel cadmium cell can be sealed as _______.
(A) it does not contain liquid electrolyte
(B) no gaseous products are formed in the reaction
(C) it is a secondary cell
(D) it is more expensive

20. A device that converts energy of combustion of fuels like hydrogen and methane, directly into electrical energy is known as _______. **[AIPMT 2015]**
(A) fuel cell (B) electrolytic cell
(C) dynamo (D) Ni-Cd cell

21. Which of the following statements is TRUE for fuel cells?
(A) They are more efficient.
(B) They are pollution free.
(C) They produce electricity as long as fuel is supplied.
(D) All of these

22. In a hydrogen-oxygen fuel cell, combustion of hydrogen occurs to _______.
(A) remove adsorbed oxygen from electrode surfaces
(B) create potential difference between the two electrodes
(C) produce high purity water
(D) generate heat

23. In $H_2 - O_2$ fuel cell, the reaction occurring at cathode is _______. **[KCET 2015]**
(A) $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$
(B) $O_{2(g)} + 2H_2O_{(l)} + 4e^- \longrightarrow 4OH^-_{(aq)}$
(C) $H^+_{(aq)} + e^- \longrightarrow \frac{1}{2}H_{2(g)}$
(D) $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

24. In a typical fuel cell, the reactant (R) and product (P) are _______. **[NEET (UG) P-II 2020]**
(A) $R = H_{2(g)}, N_{2(g)} : P = NH_{3(aq)}$
(B) $R = H_{2(g)}, O_{2(g)} : P = H_2O_{2(l)}$
(C) $R = H_{2(g)}, O_{2(g)} : P = H_2O_{(l)}$
(D) $R = H_{2(g)}, O_{2(g)}, Cl_{2(g)} : P = HClO_{4(aq)}$

25. Hydrogen-oxygen fuel cells are used in space-craft to supply _______.
(A) oxygen and hydrogen
(B) power and water
(C) CO_2 and power
(D) water and oxygen

MISCELLANEOUS

1. The anodic half-cell of lead-acid battery is recharged using electricity of 0.05 Faraday. The amount of $PbSO_4$ electrolyzed in g during the process is _______.
(Molar mass of $PbSO_4$ = 303 g mol^{-1})
[JEE (Main) Jan 2019]
(A) 7.6 (B) 15.2
(C) 11.4 (D) 22.8

2. If one end of a piece of a metal is heated, the other end becomes hot after some time. This is due to _______.
(A) energised electrons moving to the other part of the metal
(B) resistance of the metal
(C) mobility of atoms in the metal
(D) minor perturbation in the energy of atoms

3. Match the terms given in Column I with the units given in Column II.

	Column I		Column II
i.	κ	a.	S cm^2 mol^{-1}
ii.	R	b.	m^{-1}
iii.	Λ_m	c.	S cm^{-1}
iv.	G*	d.	Ω

(A) i - c, ii - d, iii - a, iv - b
(B) i - a, ii - c, iii - d, iv - b
(C) i - c, ii - b, iii - d, iv - a
(D) i - b, ii - d, iii - a, iv – c

4. Which one of the following statements is CORRECT?
(A) Nickel cadmium cell is a primary voltaic cell.
(B) The specific conductance of an electrolyte solution decreases with increase in dilution.
(C) E° is an extensive property.
(D) All of these.

5. The name of equation showing relation between electrode potential (E) standard electrode potential (E°) and concentration of ions in solution is _______.
(A) Kohlrausch's equation
(B) Nernst's equation
(C) Ohm's equation
(D) Faraday's equation

6. If same amount of electricity is passed through aqueous solutions of $AgNO_3$ and $CuSO_4$ and the number of Ag and Cu atoms deposited are x and y, respectively. Then, _______.
(A) x = y (B) x < y
(C) y = 2x (D) x = 2y

7. One litre of 1 M $CuSO_4$ solution is electrolysed. After passing 2 F of electricity, the molarity of $CuSO_4$ will be _______.
(A) M/2 (B) M/4 (C) M (D) 0

8. When electric current is supplied through an ionic hydride of fused state, then _______.
(A) hydrogen is obtained at anode
(B) hydrogen is obtained at cathode
(C) hydrogen is present in solution
(D) no change occurs

9. When electric current is passed through a cell having an electrolyte, the positive ions move towards the cathode and the negative ions towards the anode. If the cathode is pulled out of the solution, _______.
(A) the positive and negative ions will move towards the anode
(B) the positive ions will start moving towards the anode and the negative ions will stop moving
(C) the negative ions will continue to move towards the anode and the positive ions will stop moving
(D) the positive and negative ions will start moving randomly

10. Comparing a Daniel cell and a dry cell, which of the following is TRUE?
(A) Oxidation half reaction for both are same.
(B) Reduction half reaction for both are same.
(C) Oxidation half reaction for both are different.
(D) Both oxidation and reduction reactions for both cells are same.

11. In a cell that utilises the reaction,
$Zn_{(s)} + 2H^+_{(aq)} \rightleftharpoons Zn^{2+}_{(aq)} + H_{2(g)}$; addition of H_2SO_4 to cathode compartment, will _______.
(A) increase the E and shift equilibrium to the left
(B) lower the E and shift equilibrium to the right
(C) increase the E and shift equilibrium to the right
(D) lower the E and shift equilibrium to the left

12. $\underset{\text{(anode)}}{Zn_{(s)} | Zn^{2+}_{(aq)}} || \underset{\text{(cathode)}}{Cu^{2+}_{(aq)} | Cu_{(s)}}$ is _______.
(A) Daniel cell
(B) Calomel cell
(C) Faraday cell
(D) standard cell

13. Standard electrode potentials are:
Fe^{2+} | Fe, E° = – 0.44 V;
Fe^{3+} | Fe^{2+}, E° = 0.77 V.
If Fe^{2+}, Fe^{3+} and Fe blocks are kept together, it will lead to _______. **[BCECE 2015]**
(A) increase in Fe^{3+}
(B) decrease in Fe^{3+}
(C) no change in the ratio of $\frac{Fe^{2+}}{Fe^{3+}}$
(D) decrease in Fe^{2+}

14. How long (approximate) should water be electrolysed by passing through 100 amperes current so that the oxygen released can completely burn 27.66 g of diborane? (Atomic weight of B=10.8 u) **[JEE (Main) 2018]**
(A) 6.4 hours (B) 0.8 hours
(C) 3.2 hours (D) 1.6 hours

15. When 9.65 ampere current was passed for 1.0 hour into nitrobenzene in acidic medium, the amount of p-aminophenol produced is _______.
(A) 9.81 g (B) 10.9 g
(C) 98.1 g (D) 109.0 g

16. Cu^+ ion is not stable in aqueous solution because of disproportionation reaction. E° value for disproportionation of Cu^+ is ______.
(Given $E^\circ_{Cu^{2+}/Cu^+} = 0.15\,V$, $E^\circ_{Cu^{2+}/Cu} = 0.34\,V$)
(A) – 0.49 V (B) 0.49 V
(C) – 0.38 V (D) 0.38 V

Numerical Value Type Questions

1. The conductivity of 0.2 M KCl solution at 25 °C is 0.0248 S cm^{-1}. The molar conductivity of this solution is _______ S cm^2 mol^{-1}.

2. For a cell involving two electron changes, E° is 0.59 V at 25 °C. The equilibrium constant for the cell reaction is 10^x. What is the value of X?

[Given that $\frac{2.303\ RT}{F} = 0.059$ V at T = 298 K]

3. Number of Faraday's of electricity required to liberate 5.6 L of $Cl_{2(g)}$ at anode (at STP)?

4. The Gibbs change (in J) for the given reaction at $[Cu^{2+}] = [Sn^{2+}] = 1$ M and 298 K is:

$Cu_{(s)} + Sn^{2+}_{(aq.)} \longrightarrow Cu^{2+}_{(aq.)} + Sn_{(s)}$;

($E^{o}_{Sn^{2+}|Sn} = -0.16V$, $E^{o}_{Cu^{2+}|Cu} = 0.34V$,

Take F = 96500 C mol^{-1})

[JEE (Main) Sept 2020]

5. The cell potential for $Zn|Zn^{2+}(aq)||Sn^{x+}|Sn$ is 0.801 V at 298 K. The reaction quotient for the above reaction is 10^{-2}. The number of electrons involved in the given electrochemical cell reaction is _______.

(Given: $E^{0}_{Zn^{2+}|Zn} = -0.763$ V, $E^{0}_{Sn^{x+}|Sn} = +0.008$ V,

and $\frac{2.303RT}{F} = 0.06$ V)

6. At 298 K, what will be the value of electrode potential in volts for Fe^{3+}/Fe^{2+} electrode, when the concentration of Fe^{2+} is exactly 10 times that of Fe^{3+}? [Given: $E^{o}_{Fe^{3+}/Fe^{2+}} = 0.771$ V]

7. When 0.6 faraday of electricity is passed through molten $CaSO_4$, then the weight of Ca metal deposited at the cathode is W_1 g. When 0.3 faraday of electricity is passed through molten $AlCl_3$, then the weight of Al metal deposited at the cathode is W_2 g. The ratio of W_1 by W_2 is _______. (Al = 27 u, Ca = 40 u)

Topic Test

1. When a current of 3 A is passed through an aqueous solution of a palladium salt, the mass of Pd metal deposited in 1 hour is 2.98 g. If atomic mass of the metal is 106.4 u, the valency of palladium is _______.

(A) 4 (B) 5 (C) 3 (D) 2

2. The molar conductivity of NH_4Cl at infinite dilution is 149.7 S cm^2 mol^{-1} and molar conductivity of OH^- ions and Cl^- ions are 198 and 76.3 S cm^2 mol^{-1} respectively.

$\wedge^{o}_{m}$ (in S cm^2 mol^{-1}) for NH_4OH will be _______.

(A) 271.4 (B) 347.7

(C) 424.0 (D) 124.6

3. Consider the following galvanic cell:

$Cd | Cd^{2+}$ (1 M) || $H^+(aq)$ (1 M) | H_2 (g) | Pt (s)

The overall cell reaction is

(A) $Cd^{2+} + H_2 \longrightarrow Cd_{(s)} + 2H^+$

(B) $Cd^{2+} + H_2 + 2e^- \longrightarrow Cd_{(s)} + H_2\uparrow$

(C) $Cd_{(s)} + 2H^+ \longrightarrow Cd^{2+} + H_2\uparrow$

(D) $Pt + H^+ + 2e^- \longrightarrow H_2 + Pt$

4. A conductance cell was filled with a 0.02 M KCl solution which has a specific conductance of 2.768×10^{-3} ohm^{-1} cm^{-1}. If its resistance is 82.4 ohm at 25°C, the cell constant is _______.

(A) 0.2182 cm^{-1} (B) 0.2281 cm^{-1}

(C) 0.2821 cm^{-1} (D) 0.2381 cm^{-1}

5. E^{o}_{cell} for the following cell reaction is 1.75 V.

$3Ni + 2Au^{3+}(1M) \longrightarrow 3Ni^{2+}(0.1M) + 2Au$

The value of e.m.f. of the cell at 298 K is _______.

(A) 1.75 V (B) 1.78 V

(C) 1.72 V (D) 1.69 V

6. The e.m.f. of the Daniel cell,

Zn (s) | Zn^{2+}(aq) || Cu^{2+}(aq) | Cu (s) can be increased by _______.

(A) reducing the concentration of copper sulphate solution

(B) doubling concentrations of both copper sulphate and zinc sulphate solution

(C) increasing the concentration of copper sulphate solution

(D) increasing the concentration of zinc sulphate solution

7. If X is the resistivity of the solution and M is the molarity of the solution, the molar conductivity of the solution is given by _______.

(A) $\frac{1000X}{M}$ (B) $\frac{1000}{MX}$

(C) $\frac{1000M}{X}$ (D) $\frac{MX}{1000}$

8. Which of the following is NOT a strong electrolyte?
(A) Acetic acid
(B) Aqueous KOH solution
(C) Dilute HCl
(D) NaCl

9. The value of equilibrium constant for the following reaction at 298 K is _______.

$$Cr_{(s)} + Au^{3+}_{(aq)} \rightleftharpoons Au_{(s)} + Cr^{3+}_{(aq)}$$

Given that $E^\circ_{\left(Au^{3+}/Au\right)} = 1.50$ V and

$E^\circ_{\left(Cr^{3+}/Cr\right)} = -0.74$ V

(A) 10^{19} (B) 10^{38} (C) 10^{42} (D) 10^{63}

10. Which of the following is INCORRECT regarding mercury cell?
(A) It is suitable for low current devices.
(B) The electrolyte used is a paste of KOH and ZnO.
(C) The anode is made of HgO and carbon.
(D) It is a primary cell.

11. Standard electrode potentials of Zn and Fe are known to be (i) −0.76V and (ii) −0.44V, respectively. How does it explain that galvanization prevents rusting of iron while zinc slowly dissolves away?
(A) Since (i) is less than (ii), zinc becomes the cathode and iron the anode.
(B) Since (i) is less than (ii), zinc becomes the anode and iron the cathode.
(C) Since (i) is more than (ii), zinc becomes the anode and iron the cathode.
(D) Since (i) is more than (ii), zinc becomes the cathode and iron the anode.

12. The electrical resistance of a metallic conductor is _______ .
(i) directly proportional to its length
(ii) inversely proportional to its area of cross-section
(iii) inversely proportional to its length
(iv) directly proportional to its area of cross-section
(A) (i) and (ii) (B) (i) and (iv)
(C) (ii) and (iii) (D) (iii) and (iv)

13. At infinite dilution, the contribution of cation and anion to the molar conductivity is _______.
(A) dependent on each other
(B) independent of each other
(C) dependent on the nature of the solvent
(D) dependent on the cell

14. In the representation of galvanic cells, a double vertical line between two solutions indicates _______.
(A) direct contact between them
(B) that they are connected by a salt bridge
(C) the phase boundary
(D) all of these

15. Three Faradays of electricity is passed through molten $AlCl_3$ and molten NaCl taken in different electrolytic cells. The amounts of Na and Al deposited at the cathodes will be in the ratio of _______.
(A) 1 mole : 2 mole
(B) 3 mole : 2 mole
(C) 1 mole : 0.33 mole
(D) 1.5 mole : 3 mole

16. Which of the following is NOT used in making a Leclanche cell?
(A) Graphite (B) MnO_2
(C) NH_4Cl (D) PbO_2

17. Consider the standard potentials of the following half reactions:

	E° (V)
$Mg^{2+} + 2e^- \longrightarrow Mg$	–2.37
$Zn^{2+} + 2e^- \longrightarrow Zn$	–0.76
$Ni^{2+} + 2e^- \longrightarrow Ni$	–0.25
$Fe^{3+} + 3e^- \longrightarrow Fe$	+0.04

From the above list, choose the species that is the best reducing agent.
(A) Mg^{2+} (B) Mg (C) Zn (D) Fe

18. Regarding the polarity of cathode in a cell, which of the following is a TRUE statement?
(A) It is positive in an electrolytic cell and negative in a galvanic cell.
(B) It is negative in an electrolytic cell and positive in a galvanic cell.
(C) It is negative in both electrolytic and galvanic cells.
(D) It is positive in both electrolytic and galvanic cells.

19. Which reaction occurs at the anode during the electrolysis of fused lead bromide?
(A) Br^- ions are reduced.
(B) Pb^{2+} ions are oxidized.
(C) Pb^{2+} ions are reduced.
(D) Br^- ions are oxidized.

20. If the conductivity and conductance of a solution is same, then its cell constant is equal to _______ .
(A) 1 (B) 10 (C) 100 (D) 0.1

Scan the adjacent Q.R. Code in *Quill - The Padhai App* to access Answers & Solutions to Topic Test.

1 Which of the following temperature will read the same value on Celsius and Fahrenheit scales?

(A) – 40°

(B) + 40°

(C) – 80°

(D) – 20°

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