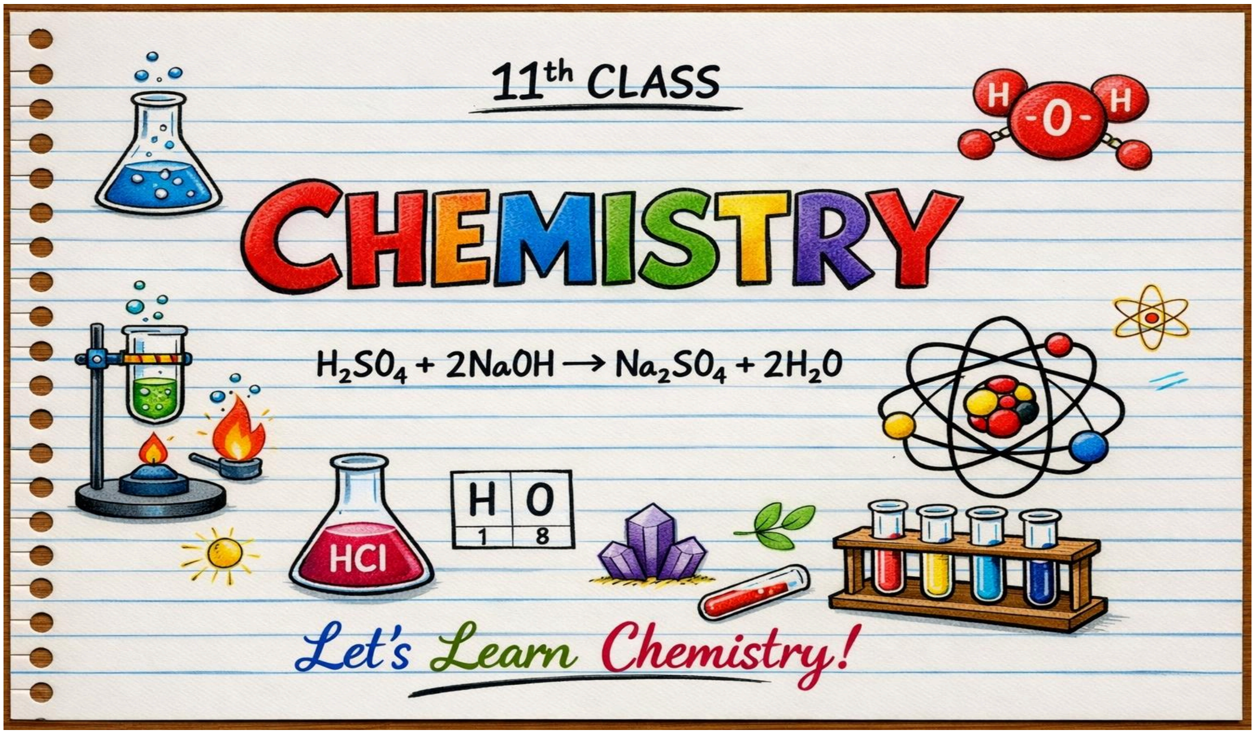

Class: 11th

Subject: Chemistry

Unit 10: ELECTROCHEMISTRY

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❖ Important MCQs:

1. Which process represents oxidation according to electron transfer concept?

(a) Gain of proton

(b) Loss of electron ✅

(c) Gain of neutron

(d) Loss of proton

2. In the reaction $Fe^{2+} \longrightarrow Fe^{3+} + e^-$, iron undergoes:

(a) Reduction

(b) Oxidation ✅

(c) Neutralization

(d) Hydrolysis

3. Which species is reduced in the following?

(a) $Zn^0 \longrightarrow Zn^{2+} + 2e^-$

(b) $Cl^0 + e^- \longrightarrow Cl^-$ ✅

(c) $Fe^{2+} \longrightarrow Fe^{3+} + e^-$

(d) $Mg \longrightarrow Mg^{2+} + 2e^-$

4. A substance that loses electrons is called:

(a) Oxidizing agent

(b) Reducing agent ✅

(c) Catalyst

(d) Acid

5. Which of the following represents reduction?

(a) Loss of electron

(b) Gain of electron ✅

(c) Loss of proton

(d) Increase in oxidation number

6. In redox reactions:

(a) Only oxidation occurs

(b) Only reduction occurs

(c) Both occur simultaneously ✅

(d) No electron transfer occurs

7. A substance that gains electrons acts as:

(a) Reducing agent

(b) Oxidizing agent ✅

(c) Base

(d) Catalyst

8. In the reaction $Zn^0 \longrightarrow Zn^{2+} + 2e^-$, zinc is:

(a) Reduced

(b) Oxidized ✅

(c) Neutral

(d) Hydrolyzed

9. Identify the oxidizing agent in the reaction:

$Cl^0 + e^- \longrightarrow Cl^-$

(a) Cl^0 ✅

(b) e^-

(c) Cl^-

(d) None

10. Oxidation can also be identified by:

(a) Decrease in oxidation number

(b) Increase in oxidation number ✅

(c) No change in oxidation number

(d) Change in physical state only

11. Oxidation number is:

(a) Real charge on atom

(b) Apparent charge on atom ✅

(c) Number of protons

(d) Atomic mass

12. Oxidation number is also called:

(a) Valency

(b) Charge

(c) Oxidation state ✅

(d) Atomic number

13. The oxidation number of an element in free state is:

(a) +1

(b) -1

(c) 0 ✅

(d) Variable

14. In a compound, the more electronegative element usually has:

(a) Positive oxidation number

(b) Negative oxidation number ✅

(c) Zero oxidation number

(d) Variable oxidation number

15. Oxidation number of Group 2 elements is:

(a) +1

(b) +2 ✅

(c) -2

(d) 0

16. Oxidation number of hydrogen in most compounds is:

(a) -1

(b) +1 ✅

(c) 0

(d) +2

17. Oxidation number of oxygen in peroxides is:

(a) -2

(b) -1 ✅

(c) +2

(d) 0

18. In OF_2, oxidation number of oxygen is:

(a) -2

(b) -1

(c) +2 ✅

(d) 0

19. Oxidation number of Cl^- ion is:

(a) 0

(b) +1

(c) -1 ✅

(d) +2

20. The sum of oxidation numbers in a neutral compound is:

(a) +1

(b) -1

(c) 0 ✅

(d) Equal to charge

21. The sum of oxidation numbers in a polyatomic ion is equal to:

(a) Zero

(b) Number of atoms

(c) Charge on ion ✅

(d) Atomic number

22. Transition elements show:

(a) Fixed oxidation states

(b) No oxidation states

(c) Variable oxidation states ✅

(d) Only zero oxidation state

23. In NH_4^+, the sum of oxidation numbers is:

(a) 0

(b) +1 ✅

(c) -1

(d) +4

24. In SO_4^{2-}, the sum of oxidation numbers is:

(a) 0

(b) +2

(c) -2 ✅

(d) +4

25. A disproportionation reaction is one in which:

(a) Only oxidation occurs

(b) Only reduction occurs

(c) Same element undergoes oxidation and reduction ✅

(d) No redox reaction occurs

26. Which statement correctly defines an oxidizing agent?

(a) It loses electrons and is reduced

(b) It gains electrons and is reduced ✅

(c) It loses electrons and is oxidized

(d) It gains electrons and is oxidized

27. In a redox reaction, the substance that is oxidized acts as:

(a) Oxidizing agent

(b) Reducing agent ✅

(c) Catalyst

(d) Neutral substance

28. Which of the following acts as a reducing agent?

(a) Substance that gains electrons

(b) Substance that loses electrons ✅

(c) Substance that gains protons

(d) Substance that gains neutrons

29. In the reaction:

$Zn \longrightarrow Zn^{2+} + 2e^-$

Zinc is acting as:

(a) Oxidizing agent

(b) Reducing agent ✅

(c) Catalyst

(d) Base

30. Which statement is correct about oxidizing agents?

(a) They are oxidized themselves

(b) They lose electrons

(c) Their oxidation number increases

(d) Their oxidation number decreases ✅

31. During balancing by oxidation number method, equalization is done for:

(a) Atoms only

(b) Charges only

(c) Change in oxidation numbers ✅

(d) Molecular mass

32. In oxidation number method, electrons lost and gained are balanced by:

(a) Adding water

(b) Multiplying coefficients ✅

(c) Changing charges

(d) Removing atoms

33. In an electrolytic cell, which electrode is positively charged?

(a) Cathode

(b) Anode ✅

(c) Both

(d) None

34. Which process occurs at cathode in an electrolytic cell?

(a) Oxidation

(b) Reduction ✅

(c) Neutralization

(d) Hydrolysis

35. Which statement best explains electrolysis?

(a) Conversion of chemical energy into heat

(b) Conversion of electrical energy into chemical energy ✅

(c) Conversion of heat into light

(d) Conversion of light into electrical energy

36. In electrolysis, the electrode where reduction occurs is:

(a) Anode

(b) Cathode ✅

(c) Electrolyte

(d) Salt bridge

37. Which species gains electrons during electrolysis?

(a) Anions

(b) Cations ✅

(c) Neutral atoms

(d) Molecules

38. In $Cu^{2+} + 2e^{-} \longrightarrow Cu(s)$, copper ions are:

(a) Oxidized

(b) Reduced ✅

(c) Neutralized

(d) Hydrolyzed

39. At anode, which process takes place?

(a) Reduction

(b) Oxidation ✅

(c) Neutralization

(d) Dissolution

40. Which statement is correct for electrolysis?

(a) Cations lose electrons

(b) Anions gain electrons

(c) Cations gain electrons at cathode ✅

(d) Anions gain electrons at anode

41. If a metal is deposited at cathode, it is due to:

(a) Oxidation

(b) Reduction ✅

(c) Neutralization

(d) Decomposition

42. Quantity of electricity is given by:

(a) Q = I/t

(b) Q = I × t ✅

(c) Q = t/I

(d) $Q = I^2t$

43. Which factor does NOT affect mass deposited during electrolysis?

(a) Current

(b) Time

(c) Quantity of charge

(d) Color of electrolyte ✅

44. One Faraday is equal to:

(a) Charge on one electron

(b) Charge on 1 mole of electrons ✅

(c) 1 coulomb

(d) 1 ampere

45. If current increases, mass deposited will:

(a) Decrease

(b) Increase ✅

(c) Remain same

(d) Become zero

46. During electrolysis of molten NaCl, product at cathode is:

(a) Cl_2

(b) Na ✅

(c) H_2

(d) O_2

47. Avogadro number is used to find:

(a) Mass of electron

(b) Number of particles in mole ✅

(c) Volume of gas

(d) Charge on ion

48. Charge on one electron is:

(a) 96500 C

(b) 6.02×10^{23} C

(c) 1.6×10^{-19} C ✅

(d) 1 C

49. Electrode potential is a measure of:

(a) Density

(b) Reactivity

(c) Ease of oxidation or reduction ✅

(d) Mass

50. In $Cu^{2+} + 2e^- \rightleftharpoons Cu$, backward reaction represents:

(a) Reduction

(b) Oxidation ✅

(c) Neutralization

(d) Combination

51. Which metal ion is more easily reduced?

(a) One with lower tendency to gain electrons

(b) One with higher tendency to gain electrons ✅

(c) One with zero charge

(d) One with higher mass

52. Electrical double layer is formed due to:

(a) Movement of atoms

(b) Separation of charges at interface ✅

(c) Temperature rise

(d) Pressure change

53. Electrode potential cannot be measured directly because of:

(a) High temperature

(b) Electrical double layer ✅

(c) Pressure

(d) Mass

54. Standard hydrogen electrode is used as:

(a) Oxidizing agent

(b) Reducing agent

(c) Reference electrode ✅

(d) Catalyst

55. Standard electrode potential is measured at:

(a) Any temperature

(b) 25°C, 1 atm, 1M concentration ✅

(c) 0°C only

(d) High pressure only

56. Standard electrode potential (E°) is measured against:

(a) Copper electrode

(b) Zinc electrode

(c) Standard hydrogen electrode (SHE) ✅

(d) Silver electrode

57. In a metal/metal ion half-cell, electrode used is:

(a) Platinum only

(b) Metal rod of same metal ✅

(c) Carbon electrode

(d) Salt bridge

58. In Cl_2/Cl^- half-cell, electrode used is:

(a) Copper

(b) Zinc

(c) Platinum or carbon electrode ✅

(d) Silver electrode

59. A more positive E° value indicates:

(a) Less tendency to gain electrons

(b) Greater tendency to gain electrons (reduction) ✅

(c) Greater tendency to lose electrons

(d) No electron transfer

60. In an electrochemical cell, electrons flow from:

(a) Cathode to anode

(b) Anode to cathode ✅

(c) Salt bridge to electrode

(d) Solution to electrode

61. In a galvanic cell, oxidation occurs at:

(a) Cathode

(b) Anode ✅

(c) Salt bridge

(d) Electrolyte only

62. In Zn/Cu galvanic cell, anode is:

(a) Copper

(b) Zinc ✅

(c) Both

(d) Salt bridge

63. In Zn/Cu cell, electrons flow from:

(a) Cu to Zn

(b) Zn to Cu ✅

(c) Salt bridge to Zn

(d) Cu to salt bridge

64. Salt bridge in a galvanic cell is used to:

(a) Transfer electrons

(b) Maintain ionic balance ✅

(c) Increase voltage

(d) Produce energy

65. Overall cell reaction in Zn–Cu cell is:

(a) $Zn + Cu \longrightarrow Zn^{2+} + Cu^{2+}$

(b) $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu(s)$ ✅

(c) $Zn^{2+} + Cu \longrightarrow Zn + Cu^{2+}$

(d) $Zn + Cu \longrightarrow ZnCu$

66. Direction of electron flow in an electrochemical cell is determined by:

(a) Atomic mass

(b) E° values of half-cells ✅

(c) Temperature

(d) Pressure

67. Electrons in external circuit flow from:

(a) Higher E° to lower E°

(b) Lower E° to higher E° ✅

(c) Cathode to anode through solution

(d) Salt bridge only

68. In a Zn–Ag cell, electron flow is from:

(a) Ag to Zn

(b) Zn to Ag ✅

(c) Ag to salt bridge

(d) Solution to electrode

69. The more positive E° value indicates species that is:

(a) Harder to reduce

(b) Easier to reduce ✅

(c) Easier to oxidize

(d) Inert

70. The more negative E° value indicates species that is:

(a) Easier to reduce

(b) Easier to oxidize ✅

(c) Neutral

(d) More stable only

71. A reaction is feasible if overall cell E° is:

(a) Negative

(b) Zero

(c) Positive ✅

(d) Infinite

72. In Mg^{2+} / Mg (E° = –2.38 V), Mg acts as:

(a) Oxidizing agent

(b) Reducing agent ✅

(c) Neutral agent

(d) Catalyst

73. In Zn^{2+} / Zn (E° = –0.76 V), Zn^{2+} is:

(a) Strong reducing agent

(b) Strong oxidizing agent ✅

(c) Neutral species

(d) Inert gas

74. A species with higher E° value is a better:

(a) Reducing agent

(b) Oxidizing agent ✅

(c) Neutral agent

(d) Catalyst

75. $KMnO_4$ acts as an oxidizing agent because it has:

(a) Low E° value

(b) High positive E° value ✅

(c) Zero E° value

(d) Negative mass

76. When concentration of reactant increases, electrode potential generally becomes:

(a) Less positive

(b) More negative

(c) More positive / less negative ✅

(d) Zero

77. Change in electrode potential under non-standard conditions is represented by:

(a) E°

(b) E only ✅

(c) F

(d) R

78. According to Le Chatelier's principle, increasing reactant concentration shifts equilibrium:

(a) Left

(b) Right ✅

(c) No change

(d) Stops reaction

79. For $Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$, standard E° = 0.77 V indicates:

(a) Fe^{3+} is harder to reduce

(b) Fe^{3+} is easier to reduce ✅

(c) Fe^{2+} is inert

(d) No reaction occurs

80. Nernst equation is used to calculate:

(a) Atomic mass

(b) Electrode potential under non-standard conditions ✅

(c) Density

(d) Pressure

81. In Nernst equation, R represents:

(a) Rate constant

(b) Gas constant ✅

(c) Resistance

(d) Radius

82. In Nernst equation, T is measured in:

(a) °C

(b) Kelvin ✅

(c) Fahrenheit

(d) Joule

83. In Nernst equation, F represents:

(a) Force

(b) Faraday constant ✅

(c) Frequency

(d) Factor

84. Number of electrons transferred in reaction is represented by:

(a) n

(b) z ✅

(c) e

(d) x

85. Nernst equation shows effect of:

(a) Pressure only

(b) Temperature and concentration ✅

(c) Mass only

(d) Volume only

86. Metals higher in activity series are:

(a) Less reactive

(b) More reactive and stronger reducing agents ✅

(c) Noble metals

(d) Inert gases

87. Alkali metals in activity series are:

(a) Least reactive

(b) Highly reactive ✅

(c) Noble metals

(d) Non-metals

88. Noble metals like gold and silver are placed:

(a) Top of activity series

(b) Middle

(c) Bottom of activity series ✅

(d) Not included

89. Metals with more negative E° values are:

(a) Less reactive

(b) More reactive and easily oxidized ✅

(c) Non-reactive

(d) Only oxidizing agents

90. Activity series of metals is based on:

(a) Atomic mass

(b) Standard reduction potentials ✅

(c) Density

(d) Color

91. A redox reaction is feasible when cell potential is:

(a) Negative

(b) Zero

(c) Positive ✅

(d) Infinite

92. Activity series of metals is arranged according to:

(a) Atomic mass

(b) Reactivity order (highest to lowest) ✅

(c) Density

(d) Valency

93. In displacement reaction, a metal will displace another if it is:

(a) Less reactive

(b) More reactive (higher in activity series) ✅

(c) Same reactivity

(d) Non-metal

94. If cell potential is negative, the reaction is:

(a) Spontaneous

(b) Feasible

(c) Not feasible ✅

(d) Neutral

95. Photovoltaic cell converts:

(a) Chemical energy into heat

(b) Light energy into electrical energy ✅

(c) Electrical energy into chemical energy

(d) Mechanical energy into light

96. Photovoltaic effect involves excitation of electrons from:

(a) Conduction band to valence band

(b) Valence band to conduction band ✅

(c) Nucleus to orbit

(d) Proton to neutron

97. PN junction in photovoltaic cell is used to:

(a) Increase resistance

(b) Separate charges and generate voltage ✅

(c) Stop current flow

(d) Produce heat only

98. Photovoltaic systems are sustainable because they:

(a) Use coal

(b) Depend on fossil fuels

(c) Use sunlight as energy source ✅

(d) Produce pollution

99. BOD measures:

(a) Oxygen in air

(b) Oxygen required for decomposition of organic matter by bacteria ✅

(c) Carbon dioxide in water

(d) Salt concentration

100. Dissolved Oxygen (DO) refers to:

(a) Oxygen in soil

(b) Oxygen dissolved in water ✅

(c) Oxygen in rocks

(d) Oxygen in plants only

101. DO level below 5 mg/dm^3 indicates:

(a) Pure water

(b) Polluted water ✅

(c) Distilled water

(d) Oxygen rich water

102. BOD and DO have relationship:

(a) Direct

(b) No relation

(c) Inverse relationship ✅

(d) Equal values

103. In Winkler method, DO is measured by:

(a) Weighing method

(b) Iodine titration method ✅

(c) Heating method

(d) Filtration method

104. In Winkler method, iodine is titrated with:

(a) HCl

(b) NaOH

(c) Sodium thiosulfate ✅

(d) $KMnO_4$

105. Starch is used in Winkler method as:

(a) Acid

(b) Indicator ✅

(c) Catalyst

(d) Oxidizing agent

106. Blue color in Winkler method indicates:

(a) End point reached ✅

(b) No reaction

(c) Excess oxygen

(d) No iodine present

107. Activity series helps to predict:

(a) Color change

(b) Reaction feasibility and displacement reactions ✅

(c) Temperature change

(d) Mass change

108. Metals at top of activity series are:

(a) Least reactive

(b) Most reactive and strong reducing agents ✅

(c) Noble metals

(d) Inert gases

109. Photovoltaic effect depends on:

(a) Heat energy

(b) Photon energy exciting electrons ✅

(c) Pressure

(d) Magnetic field only

110. In PV cell, current is produced due to:

(a) Movement of protons

(b) Movement of electrons and holes separation ✅

(c) Movement of neutrons

(d) Chemical burning

❖ Important Short Questions:

1. Define oxidation.

Answer:

- Oxidation is a chemical process in which a substance loses electrons during a reaction.
- It is also defined as an increase in the oxidation number of an element.
- Oxidation always occurs along with reduction in redox reactions.
- During oxidation, the species becomes more positively charged or less negatively charged.

Example: $Zn \rightarrow Zn^{2+} + 2e^-$, here zinc loses electrons and is oxidized.

2. Define reduction.

Answer:

- Reduction is a chemical process in which a substance gains electrons during a reaction.

- It is also defined as a decrease in the oxidation number of an element.
- Reduction always occurs simultaneously with oxidation in redox reactions.
- During reduction, the species becomes less positive or more negative.

Example: $Cl_2 + 2e^- \longrightarrow 2Cl^-$, chlorine gains electrons and is reduced.

3. What is a redox reaction?

Answer:

- A redox reaction is a chemical reaction in which oxidation and reduction occur at the same time.
- In this reaction, one substance loses electrons while another gains electrons.
- The electron transfer between species is the main feature of redox reactions.
- Both processes are always linked and cannot occur separately.

Example: $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$.

4. Define oxidation number.

Answer:

- Oxidation number is the apparent charge assigned to an atom in a compound or ion.
- It represents the degree of oxidation or reduction of an element.

- It can be positive, negative, or zero depending on the element.
- It helps in identifying electron transfer in redox reactions.

Example: In NaCl, Na = +1 and Cl = -1.

5. What is the oxidation number of a free element?

Answer:

- The oxidation number of any free or uncombined element is always zero.
- This is because the element is not bonded with any other element.
- In its elemental form, there is no charge on the atoms.
- This rule applies to all elements in their natural state.

Example: H_2, O_2, Cl_2, Fe all have oxidation number = 0.

6. What is the oxidation number of oxygen in most compounds?

Answer:

- The oxidation number of oxygen is usually -2 in most compounds.
- This is because oxygen is highly electronegative and gains electrons.
- It helps in balancing charges in compounds.
- There are exceptions like peroxides where it is -1 and OF_2 where it is +2.

Example: In H_2O, oxygen has oxidation number -2.

7. What is the oxidation number of hydrogen in compounds?

Answer:

- The oxidation number of hydrogen is usually +1 in most compounds.
- This is because hydrogen usually loses one electron in bonding.
- In metal hydrides, hydrogen shows oxidation number -1.
- This exception occurs when hydrogen is bonded with metals.

Example: In HCl, H = +1; in NaH, H = -1.

8. Define oxidizing agent.

Answer:

- An oxidizing agent is a substance that gains electrons in a chemical reaction.
- It causes oxidation of another substance and itself gets reduced.
- It increases its oxidation state is reduced during the reaction.
- It is responsible for electron acceptance in redox reactions.

Example: $Cl_2 + 2e^- \rightarrow 2Cl^-$, chlorine is the oxidizing agent.

9. Define reducing agent.

Answer:

- A reducing agent is a substance that loses electrons in a chemical reaction.
- It causes reduction of another substance and itself gets oxidized.

- Its oxidation number increases during the reaction.
- It donates electrons in redox reactions.

Example: Zn ⟶ Zn^{2+} + $2e^-$, zinc is the reducing agent.

10. At which electrode does oxidation occur?

Answer:

- Oxidation always occurs at the anode in electrochemical cells.
- At the anode, the substance loses electrons.
- These electrons move through the external circuit to the cathode.
- This is a key rule in both electrolytic and galvanic cells.

Example: Zn ⟶ Zn^{2+} + $2e^-$ occurs at the anode.

11. At which electrode does reduction occur?

Answer:

Reduction occurs at the cathode in an electrochemical cell. At this electrode, substances gain electrons. This decreases their oxidation number. For example, Cu^{2+} + $2e^-$ ⟶ Cu occurs at the cathode.

12. What is electrode potential?

Answer:

Electrode potential is the tendency of an electrode to gain or lose electrons. It shows how easily a substance is reduced or oxidized. It is measured in volts. It helps to determine the direction of electron flow.

13. What is standard hydrogen electrode (SHE)?

Answer:

Standard hydrogen electrode is a reference electrode with a potential of 0.00 V. It consists of hydrogen gas at 1 atm and 1 M H^+ ions. A platinum electrode is used. It is used to measure other electrode potentials.

14. What is a galvanic cell?

Answer:

A galvanic cell converts chemical energy into electrical energy. It works through a spontaneous redox reaction. Oxidation occurs at anode and reduction at cathode. Electrons flow through an external circuit.

15. What is the function of a salt bridge?

Answer:

A salt bridge maintains electrical neutrality in a cell. It allows movement of ions between half-cells. It completes the circuit and prevents charge buildup. It usually contains an inert electrolyte.

16. Write the formula for electrical charge in electrolysis.

Answer:

The formula for electrical charge is $Q = I \times t$. Here Q is charge, I is current and t is time. It shows that charge depends on current and time. It is used in electrolysis calculations.

17. What is one Faraday?

Answer:

One Faraday is the charge of one mole of electrons. Its value is 96500 coulombs. It is used in electrolysis calculations. It represents Avogadro number of electrons.

18. Define photovoltaic cell.

Answer:

A photovoltaic cell converts light energy into electrical energy. It works on photovoltaic effect. Light excites electrons to produce current. It uses a PN junction.

19. What is BOD?

Answer:

BOD is the amount of oxygen required to decompose organic matter in water. It is measured over five days. High BOD shows polluted water. It indicates water quality.

20. What is dissolved oxygen (DO)?

Answer:

DO is the amount of oxygen dissolved in water. It is necessary for aquatic life. It comes from air or photosynthesis. Low DO indicates pollution.

21. What is Winkler method used for?

Answer:

Winkler method is used to measure dissolved oxygen in water. It involves iodine formation and titration. Sodium thiosulfate is used. Starch acts as an indicator.

22. What is activity series of metals?

Answer:

Activity series is the arrangement of metals based on reactivity. More reactive metals are at the top. It helps to predict displacement reactions. It is also called reactivity series.

23. What is meant by feasibility of a reaction?

Answer:

Feasibility means whether a reaction can occur or not. It depends on cell potential. Positive value means reaction is possible. Negative means not feasible.

24. What is Nernst equation used for?

Answer:

Nernst equation is used to calculate electrode potential in non-standard conditions. It relates potential with concentration and temperature. It gives accurate results. It is important in electrochemistry.

25. What is electrolysis?

Answer:

Electrolysis is the process of using electrical energy to drive a chemical reaction. It occurs in electrolytic cells. Reduction occurs at cathode and oxidation at anode. It is used in metal extraction.

❖ Important Long Questions:

🌟 Q1. Define oxidation and reduction. Explain redox reactions with examples.

❖ Answer:

Oxidation and reduction are important chemical processes that always occur together in a reaction.

Oxidation is defined as the process in which a substance loses electrons. It can also be described as an increase in oxidation number. For example:

$Zn \rightarrow Zn^{2+} + 2e^-$

In this reaction, zinc loses electrons, so it is oxidized.

Reduction is defined as the process in which a substance gains electrons. It can also be described as a decrease in oxidation number.

For example:

$Cl_2 + 2e^- \longrightarrow 2Cl^-$

Here, chlorine gains electrons, so it is reduced.

A redox reaction is a chemical reaction in which oxidation and reduction occur simultaneously. In such reactions, one substance loses electrons while another gains electrons. These processes cannot occur independently; both must happen together.

For example:

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

In this reaction, zinc loses electrons and is oxidized, while copper ions gain electrons and are reduced.

Thus, redox reactions involve the transfer of electrons between substances and play an important role in many natural and industrial processes such as respiration, corrosion, and electrolysis.

🌟 **Q2. Explain oxidation number. Discuss its rules with suitable examples.**

❖ **Answer:**

Definition of Oxidation Number

Oxidation number (oxidation state) is the apparent charge assigned to an atom in a compound or ion. It shows the degree of oxidation or reduction of an element. It may be positive, negative, or zero depending on the bonding. It is helpful in identifying redox reactions and balancing equations.

Rules for Assigning Oxidation Number

1. Oxidation number of free elements

The oxidation number of any uncombined element is always zero.

Example: H_2, O_2, Na ⟶ 0

2. More electronegative element

In a compound, the more electronegative element gets a negative oxidation number.

Example: In HF, F = −1

3. Fixed oxidation numbers

Group 1 elements = +1, Group 2 = +2

Hydrogen = +1 (except hydrides)

Oxygen = −2 (except peroxides)

Example: H_2O ⟶ H = +1, O = −2

4. Monoatomic ions

The oxidation number of a monoatomic ion is equal to its charge.

Example: Na^+ = +1, Cl^- = −1

5. Sum in compounds

The sum of oxidation numbers in a neutral compound is zero.

Example: NaCl ⟶ (+1) + (−1) = 0

6. Sum in polyatomic ions

The sum of oxidation numbers equals the charge on the ion.

Example: SO_4^{2-} = −2

Example Explanation

In H_2O, hydrogen has +1 oxidation number and oxygen has −2. The total becomes zero, which satisfies the rule. This shows how oxidation numbers are assigned correctly.

Summary:

Oxidation number is an important concept in chemistry. It helps to identify oxidation and reduction processes. It is also useful in balancing redox equations and understanding chemical reactions.

🌟 Q3. Define oxidizing agent and reducing agent. Explain with examples.

❖ **Answer:**

Definition of Oxidizing Agent

An oxidizing agent is a substance that gains electrons during a chemical reaction. It causes oxidation of another substance and itself gets reduced. Its oxidation number decreases as it accepts electrons. Oxidizing agents play an important role in redox reactions.

Example of Oxidizing Agent

$Cl_2 + 2e^- \longrightarrow 2Cl^-$

In this reaction, chlorine gains electrons and is reduced. Therefore, Cl_2 acts as an oxidizing agent because it causes oxidation of another substance.

Definition of Reducing Agent

A reducing agent is a substance that loses electrons during a chemical reaction. It causes reduction of another substance and itself gets oxidized. Its oxidation number increases as it donates electrons. Reducing agents are electron donors.

Example of Reducing Agent

$Zn \longrightarrow Zn^{2+} + 2e^-$

Here, zinc loses electrons and is oxidized. Therefore, Zn acts as a reducing agent because it provides electrons to another substance.

Explanation in a Redox Reaction

In the reaction:

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

Zinc loses electrons and acts as a reducing agent, while Cu^{2+} gains electrons and acts as an oxidizing agent. This shows that both agents work together in a redox reaction.

Summary:

Oxidizing and reducing agents are essential parts of redox reactions. One accepts electrons while the other donates electrons. Their identification helps in understanding electron transfer and chemical processes.

🌟 Q4. Describe the construction and working of an electrolytic cell.

❖ Introduction

An electrolytic cell is a device in which electrical energy is converted into chemical energy through a non-spontaneous redox reaction. It is used in electrolysis to decompose compounds into their elements or simpler substances.

Construction of an Electrolytic Cell

An electrolytic cell consists of a container filled with an electrolyte, which may be a molten ionic compound or an aqueous solution of ions. Two electrodes, called anode and cathode, are immersed in the electrolyte. These electrodes are usually made of inert materials like graphite or platinum, or sometimes reactive metals depending on the process. The electrodes are connected to an external direct current (DC) power supply. The cathode is connected to the negative terminal of the battery, while the anode is connected to the positive terminal. A complete circuit is formed for the flow of electricity.

Working of an Electrolytic Cell

When electric current is passed through the electrolyte, it dissociates into positive and negative ions. The positive ions (cations) move towards the cathode where they gain electrons and undergo reduction. The negative ions (anions) move towards the anode where they lose electrons and undergo oxidation. Thus, reduction always occurs at the cathode and oxidation always occurs at the anode. The movement of ions and transfer of electrons produce chemical changes in the electrolyte.

Example Reactions

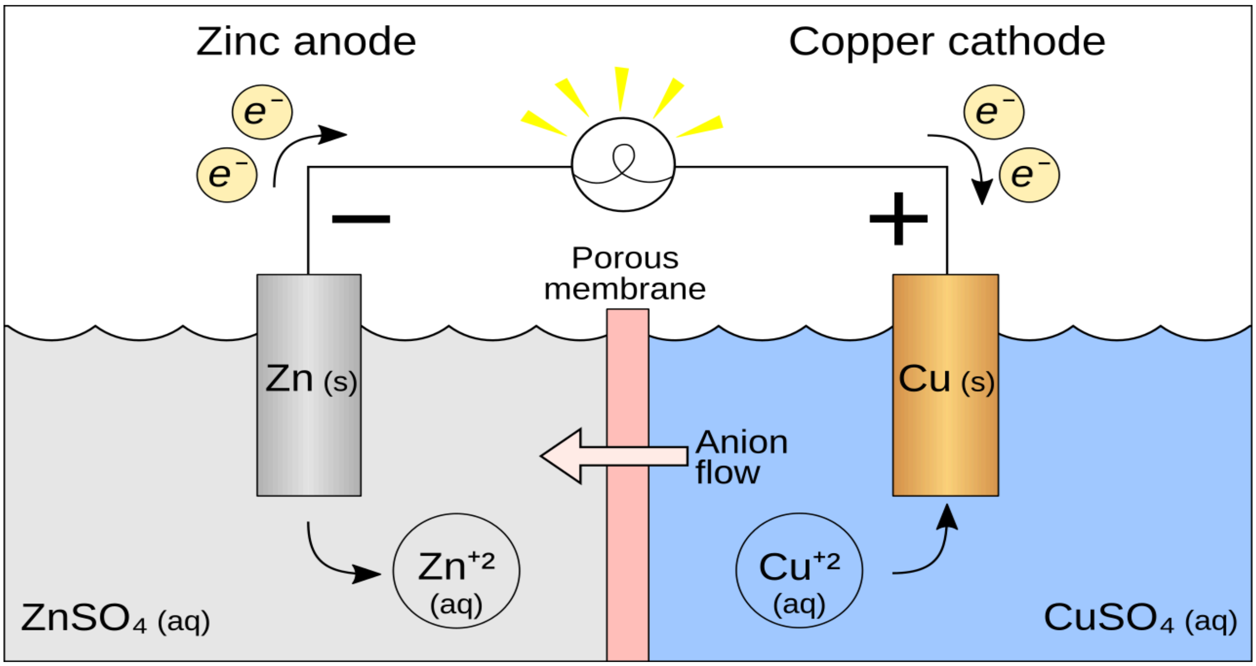
At cathode: $Cu^{2+} + 2e^- \longrightarrow Cu$ (reduction)

At anode: $2Cl^- \longrightarrow Cl_2 + 2e^-$ (oxidation)

Summary:

An electrolytic cell works by using electrical energy to drive a chemical reaction. It involves the movement of ions and electron transfer at electrodes. It is widely used in electrolysis, metal extraction, and electroplating processes.

🌟 Q5. Explain the galvanic (electrochemical) cell with diagram and working.



❖ Introduction

A galvanic cell (also called an electrochemical cell) is a device in which chemical energy is converted into electrical energy through a spontaneous redox reaction. In this cell, oxidation and reduction reactions occur in separate half-cells, and electrons flow through an external circuit.

Construction of Galvanic Cell

A galvanic cell consists of two half-cells. Each half-cell contains a metal electrode dipped in a solution of its ions. For example, in a Zn–Cu cell, zinc metal is dipped in Zn^{2+} solution and copper metal is dipped in Cu^{2+} solution. The two electrodes are connected through an external wire and a voltmeter. A salt bridge containing an inert electrolyte like KNO_3 connects the two solutions to maintain electrical neutrality and complete the circuit.

Working of Galvanic Cell

In a galvanic cell, oxidation takes place at the anode and reduction takes place at the cathode. In a Zn–Cu cell, zinc acts as the anode and undergoes oxidation:

$Zn \longrightarrow Zn^{2+} + 2e^-$

The released electrons travel through the external circuit to the copper electrode. At the cathode, copper ions gain electrons and are reduced:

$Cu^{2+} + 2e^- \longrightarrow Cu$

This flow of electrons produces electric current which can be used to do electrical work.

Role of Salt Bridge

The salt bridge maintains electrical neutrality by allowing the movement of ions between the two half-cells. It prevents charge buildup and completes the internal circuit.

Summary:

A galvanic cell converts chemical energy into electrical energy through a spontaneous redox reaction. It involves electron flow from anode to cathode and is widely used in batteries and electrochemical devices.

📘 Exercise

Q.1 Four choices are given for each question. Select the correct choice.

I. The activity series of metals arranges metals in order of their:

(a) Atomic mass

(b) Density

(c) Ease of oxidation ✅

(d) Ease of reduction

II. According to the activity series, which of the following metals would most readily displace hydrogen gas from dilute acids?

(a) Copper (Cu)

(b) Silver (Ag)

(c) Magnesium ✅

(d) Platinum (Pt)

III. The salt bridge allows transfer of in Zn-Cu voltaic cell.

(a) Zn^{2+} ions

(b) SO_4^{2-} ions

(c) Both ✅

(d) None of these

IV. If Zn-Cu galvanic cell works ideally after complete discharge, both compartments will have:

(a) $CuSO_4$ solution

(b) $ZnSO_4$ solution ✅

(c) Cu^{2+} ions

(d) Zn solid

V. Which of the following half-reactions represents a reduction process?

(a) $Zn \rightarrow Zn^{2+} + 2e^-$

(b) $Fe^{2+} \rightarrow Fe^{3+} + e^-$

(c) $Cl_2 + 2e^- \rightarrow 2Cl^-$ ✅

(d) $2H^+ + 2e^- \rightarrow H_2$

VI. Which of the following half-reactions represents the oxidation process occurring in the disproportionation of Cu^+?

(a) $Cu^{2+} + 2e^- \rightarrow Cu$

(b) $Cu^+ + e^- \rightarrow Cu$

(c) $Cu^+ \rightarrow Cu^{2+} + e^-$ ✅

(d) $2Cu^+ \rightarrow Cu^{2+} + Cu$

VII. If salt bridge is not used between two half cells in a Galvanic cell, then the voltage

(a) Decrease slowly

(b) Decrease rapidly

(c) Does not change

(d) Drop to zero ✅

VIII. In an electrolysis experiment, if a charge of 96,500 Coulombs is passed through a solution, the amount of substance liberated or deposited at the electrode is directly related to:

(a) Mass number of the ion

(b) One mole of electrons being transferred ✅

(c) Avogadro's number of ions being discharged

(d) Standard electrode potential of the metal ion

IX. The experimental determination of Avogadro's number through electrolysis typically involves measuring:

(a) The current and voltage applied

(b) The mass of the substance deposited or liberated by a known charge ✅

(c) The conductivity of the electrolytic solution

(d) The temperature changes during electrolysis

X. The principle of measuring DO by Winkler's Method is based on

(a) Iodimetry

(b) Iodometry ✅

(c) Acid-Base titration

(d) Complexometry

XI. A positive value for the standard electrode potential (E°) indicates that:

(a) The metal is a strong reducing agent

(b) The metal ion is readily oxidized

(c) The metal ion is readily reduced ✅

(d) The metal will readily displace hydrogen

XII. Which of the following changes would typically lead to an increase in the rate of electrolysis?

(a) Decreasing the concentration

(b) Increasing distance between electrodes

(c) Decreasing surface area

(d) Increasing the current passed through the cell ✅

Q.2 Attempt the following short-answer questions:

a. How and why is the electrical double layer formed?

When a metal electrode is dipped into its electrolyte, some metal atoms lose electrons and enter the solution as ions. This leaves excess electrons on the electrode, making it negatively charged, while the solution near the electrode becomes positively charged. Thus, two layers of opposite charges are formed, called the electrical double layer.

b. Why is the electrode potential of Cu called reduction potential?

The electrode potential of copper is called reduction potential because the tendency of Cu^{2+} ions to gain electrons and deposit as copper metal is measured:

$Cu^{2+} + 2e^{-} \longrightarrow Cu$

Hence, it represents reduction.

c. What are the advantages of salt bridge in a galvanic cell?

Salt bridge:

- Maintains electrical neutrality by migration of ions
- Completes the electrical circuit
- Prevents direct mixing of electrolytes
- Minimizes liquid junction potential

d. How can we predict the feasibility of a reaction using cell voltage?

The feasibility of a reaction is determined by the value of E°cell.

- If E°cell is positive, the reaction is spontaneous and feasible.
- If E°cell is negative, the reaction is non-spontaneous and not feasible.

e. During electrolysis of aqueous NaCl, why is Na not liberated at the cathode?

In aqueous NaCl, both Na^+ and H_2O are present. Water has a lower reduction potential than Na^+, so it is reduced preferentially, producing hydrogen gas:

$2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$

Therefore, sodium metal is not deposited.

f. Calculate the oxidation number of chromium (Cr) in the following compounds:

(i) $CrCl_3$

Cl = -1

Let oxidation number of Cr = x

x + 3(-1) = 0

x - 3 = 0

x = +3

✔ Oxidation number of Cr = +3

(ii) $Cr_2(SO_4)_3$

SO_4 = -2

3 SO_4 = -6

Let Cr = x

2x + (-6) = 0

2x = +6

x = +3

✔ Oxidation number of Cr = +3

(iii) $Cr_2O_7^{2-}$

O = -2

7 O = -14

Let Cr = x

$2x - 14 = -2$

$2x = +12$

$x = +6$

✔ Oxidation number of Cr = +6

g. The order of decreasing reactivity of metals is:

K > Mg > Zn > Fe > Cu

Write balanced chemical equations:

(i) Copper is added to magnesium sulfate solution ($MgSO_4$)

Copper is less reactive than magnesium, so it cannot displace Mg.

✔ No reaction occurs

(ii) Iron is added to dilute hydrochloric acid (HCl)

Iron is more reactive than hydrogen, so it displaces hydrogen gas:

$Fe + 2HCl \longrightarrow FeCl_2 + H_2$⬆️

✔ Balanced equation is: $Fe + 2HCl \longrightarrow FeCl_2 + H_2$

h. Explain why some metals higher in the activity series can displace hydrogen from acids, while others lower cannot.

Metals higher in the activity series have a greater tendency to lose electrons (they are more reactive).

They can easily form positive ions and displace hydrogen ions (H^+) from acids to produce hydrogen gas.

Metals lower in the activity series do not lose electrons easily, so they cannot displace hydrogen from acids.

i. Calculate the number of Faradays required to deposit:

✔ Basic rule: 1 Faraday = 1 mole of electrons

(i) 108 g of Ag^+

Atomic mass of Ag = 108 g/mol

Ag^+ needs 1 electron

✔ 1 Faraday required

(ii) 63.5 g of Cu^{2+}

Atomic mass of Cu = 63.5 g/mol

Cu^{2+} needs 2 electrons

✔ 2 Faradays required

(iii) 27 g of Al^{3+}

Atomic mass of Al = 27 g/mol

Al^{3+} needs 3 electrons

✔ 3 Faradays required

j. In an electrolysis experiment, a current of 0.500 A was passed through a solution of $AgNO_3$ for 30.0 minutes. The mass of silver deposited on the cathode was found to be 0.503 g. Given that the molar mass of silver is 107.87 g mol^{-1} and the charge on a silver ion is +1. Calculate the value of Avogadro's number (N_a) from this data.

Solution:

Given:

I = 0.500 A

t = 30.0 min = 30 × 60 = 1800 s

Mass of Ag = 0.503 g

Molar mass of Ag = 107.87 g mol^{-1}

Valency of Ag^+ = 1

Step 1: Total charge passed

Q = I × t

Q = 0.500 × 1800 = 900 C

Step 2: Moles of Ag deposited

Moles = mass / molar mass

= 0.503 / 107.87

= 4.66×10^{-3} mol

Step 3: Charge for 1 mole electrons (Faraday constant)

For Ag^+ (1 electron),

900 C corresponds to 4.66×10^{-3} mol

So,

$F = 900 / (4.66 \times 10^{-3})$

$F = 1.93 \times 10^5$ C mol^{-1}

Step 4: Calculate Avogadro's number

$F = N_a \times e$

$e = 1.6 \times 10^{-19}$ C

$N_a = F / e$

$N_a = (1.93 \times 10^5) / (1.6 \times 10^{-19})$

✔ $N_a = 1.21 \times 10^{24}$ mol^{-1}

k. A cell is set up with a standard nickel electrode and a standard cobalt electrode:

$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$ $E° = −0.25$ V

$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$ $E° = −0.28$ V

(i) Identify which metal will be the anode and which will be the cathode. Justify your answer.

In a galvanic cell, the electrode with more negative standard reduction potential (E°) undergoes oxidation and acts as the anode, while the one with less negative E° undergoes reduction and acts as the cathode.

Co^{2+}/Co = −0.28 V (more negative) ⟶ Anode (oxidation occurs)

Ni^{2+}/Ni = −0.25 V (less negative) ⟶ Cathode (reduction occurs)

✔ Anode: Cobalt (Co)

✔ Cathode: Nickel (Ni)

(ii) Write the balanced overall cell reaction.

Anode reaction (oxidation):

$Co(s) \longrightarrow Co^{2+}(aq) + 2e^-$

Cathode reaction (reduction):

$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$

Adding both:

✔ Overall reaction:

$Co(s) + Ni^{2+}(aq) \longrightarrow Co^{2+}(aq) + Ni(s)$

(iii) Calculate the standard cell potential (E°cell).

Formula:

$E°_{cell} = E°_{cathode} - E°_{anode}$

= (−0.25 V) − (−0.28 V)

= −0.25 + 0.28

= +0.03 V

✔ **Final Answer:** E°cell = 0.03 V

❖ DESCRIPTIVE QUESTIONS

🌟 **Q.3 How electrode potential varies with concentration of an aqueous solution? Use the Nernst equation to explain this variation.**

❖ **Answer:**

The electrode potential of a half-cell depends on the concentration of ions present in the solution. When the concentration changes, the position of equilibrium shifts according to Le-Chatelier's principle, which causes a change in electrode potential. This relationship between electrode potential and concentration is explained by the Nernst equation.

The Nernst equation is written as:

$E = E° − (RT / nF) \ln ([reduced] / [oxidized])$

At 25°C, it is simplified as:

$E = E° − (0.059 / n) \log ([reduced] / [oxidized])$

According to this equation, if the concentration of the oxidized species increases, the electrode potential becomes more positive, indicating a greater tendency for reduction. On the other hand, if the concentration of the reduced species increases, the electrode potential becomes less positive or more negative, showing a lower tendency for reduction.

For example, in the reaction $Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$, increasing the concentration of Fe^{3+} ions increases the value of electrode potential, while increasing the concentration of Fe^{2+} ions decreases it.

Thus, the electrode potential varies with concentration, and the Nernst equation provides a quantitative way to calculate this variation under non-standard conditions.

🌟 Q.4. How can Avogadro's number be derived using an electrolytic cell?

❖ Introduction

Avogadro's number is the number of particles (atoms, molecules, or ions) present in one mole of a substance. It can be determined experimentally using an electrolytic cell by measuring the charge required to deposit a known amount of substance during electrolysis.

Principle of Method

The method is based on the fact that electrolysis involves the transfer of electrons. By measuring the total charge passed and the amount of substance deposited, we can calculate the charge required for one mole of electrons, known as one Faraday.

Experimental Process

In an electrolytic cell, a constant electric current is passed through an electrolyte for a known time. The total charge is calculated using the formula $Q = I \times t$. During this process, a substance is deposited at the electrode, and its mass is measured. From the mass, the number of moles of the substance is calculated.

Relation with Faraday Constant

From the experiment, the charge required to deposit one mole of a substance is determined. This charge is called one Faraday and has a value of 96500 coulombs. It represents the charge of one mole of electrons.

Calculation of Avogadro's Number

The charge on a single electron is known to be 1.6×10^{-19} coulombs. Avogadro's number is calculated by dividing the total charge of one mole of electrons by the charge of one electron:

$N_a = 96500 / (1.6 \times 10^{-19}) \approx 6.02 \times 10^{23}\ mol^{-1}$

Summary:

Thus, by using an electrolytic cell and measuring the charge passed during electrolysis, Avogadro's number can be determined. This method provides a practical way to relate macroscopic measurements with microscopic particles.

🌟 Q.5. Describe the construction and working principle of the Zn–Cu galvanic cell.

❖ Introduction

A Zn–Cu galvanic cell is an electrochemical cell that converts chemical energy into electrical energy through a spontaneous redox reaction. It consists of two half-cells: one containing zinc and the other containing copper.

Construction of Zn–Cu Cell

The cell consists of two separate containers. In one container, a zinc rod is dipped in a solution of zinc sulfate ($ZnSO_4$), and in the other, a copper rod is dipped in a solution of copper sulfate ($CuSO_4$). The two electrodes are connected externally by a wire and a voltmeter. The two solutions are connected internally by a salt bridge containing an inert electrolyte like potassium nitrate (KNO_3), which maintains electrical neutrality.

Working Principle

The working of the Zn–Cu cell is based on a redox reaction. Zinc, being more reactive, undergoes oxidation at the anode by losing electrons:

$Zn \longrightarrow Zn^{2+} + 2e^-$

The electrons flow through the external circuit from the zinc electrode to the copper electrode. At the copper electrode (cathode), copper ions gain electrons and are reduced:

$Cu^{2+} + 2e^{-} \longrightarrow Cu$

This movement of electrons produces an electric current in the circuit.

Overall Cell Reaction

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

Role of Salt Bridge

The salt bridge allows the movement of ions between the two half-cells to maintain electrical neutrality. It prevents the accumulation of charges and ensures the continuous flow of current.

Summary:

The Zn–Cu galvanic cell works on the principle of spontaneous redox reaction, where oxidation occurs at the anode and reduction at the cathode. It produces electrical energy and is a basic model for understanding batteries.

🌟 Q.6. What is meant by Standard Hydrogen Electrode (SHE)? How is it used to measure the electrode potential of another electrode?

❖ **Definition of SHE**

The Standard Hydrogen Electrode (SHE) is a reference half-cell used to measure the electrode potential of other electrodes. It consists of hydrogen gas at 1 atm pressure in contact with 1.0 mol dm^{-3} H^+ ions solution. A platinum electrode coated with platinum black is used as an inert surface. The standard electrode potential of SHE is taken as 0.00 V.

Construction of SHE

Hydrogen gas is continuously bubbled over a platinum electrode dipped in an acidic solution containing H^+ ions. The platinum does not take part in the reaction; it only provides a surface for electron transfer. The half-reaction is:

$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$

Measurement of Electrode Potential

To measure the electrode potential of another electrode, the unknown half-cell is connected with SHE to form a complete electrochemical cell. The voltmeter measures the potential difference between the two electrodes. This value gives the standard electrode potential (E°) of the unknown electrode.

If electrons flow from SHE to the other electrode, then that electrode has a higher tendency to gain electrons (positive E°). If electrons flow from the other electrode to SHE, then that electrode has a negative E° value.

Example

When a copper electrode is connected with SHE, the measured voltage is +0.34 V. This means Cu^{2+}/Cu has a higher reduction tendency than hydrogen, so Cu acts as the cathode.

Summary:

Thus, SHE acts as a universal reference electrode with zero potential. It is used to determine the standard electrode potential of all other half-cells by comparing voltage differences in an electrochemical cell.