CHEMISTRY

Unit 10

# ELECTROCHEMISTRY

## MCQs • Short Questions • Long Questions

Complete Study Booklet with Answer Keys

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# Section 1 | Multiple Choice Questions

Tick the correct option. Correct answers are highlighted in green with a check mark.

**1. The activity series of metals arranges metals in order of their:**

(a) Atomic mass
(b) Density
✓ **(c) Ease of oxidation**
(d) Ease of reduction

**2. Which of the following metals would most readily displace hydrogen gas from dilute acids?**

(a) Copper (Cu)
(b) Silver (Ag)
✓ **(c) Magnesium (Mg)**
(d) Platinum (Pt)

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

(a) Zn2+ ions
✓ **(b) SO42- ions**
(c) Both
(d) None of these

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

(a) CuSO4 solution
✓ **(b) ZnSO4 solution**
(c) Cu2+ ions
(d) Zn solid

**5. Which of the following half-reaction represents a reduction process?**

(a) Zn(s) -> Zn2+(aq) + 2e-
(b) Fe2+(aq) -> Fe3+(aq) + e-
✓ **(c) Cl2(g) + 2e- -> 2Cl-(aq)**
(d) 2H+(aq) + 2e- -> H2(g)

**6. Which of the following half-reaction represents the oxidation process occurring in the disproportionation of Cu+?**

(a) Cu2+(aq) + 2e- -> Cu(s)
(b) Cu+(aq) + e- -> Cu(s)
✓ **(c) Cu+(aq) -> Cu2+(aq) + e-**
(d) 2Cu+(aq) -> Cu2+(aq) + Cu(s)

**7. 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**

**8. 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

**9. 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

**10. The principle of measuring DO by Winkler's Method is based on:**

(a) Iodimetry
✓ **(b) Iodometry**
(c) Acid-Base titration
(d) Complexometry

**11. A positive value for the standard electrode potential (E deg ) of a metal ion/metal half-cell (e.g., Cu2+/Cu) indicates that:**

(a) The metal is strong reducing agent
(b) The metal ion is readily oxidized
✓ **(c) The metal ion is readily reduced**
(d) The metal will readily displace hydrogen from dilute acids

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

(a) Decreasing the concentration of the electrolyte
(b) Increasing the distance between the electrodes
(c) Decreasing the surface area of the electrolytic cell
✓ **(d) Increasing the current passed through the electrolytic cell**

**13. Which of the following best defines oxidation?**

(a) Gain of electrons
(b) Loss of protons
✓ **(c) Loss of electrons**
(d) Gain of neutrons

**14. What is the oxidation number of chlorine in NaClO3?**

(a) +3
✓ **(b) +5**
(c) +1
(d) -1

**15. In the reaction: Fe -> Fe2+ + 2e-, iron is:**

(a) Reduced
✓ **(b) Oxidized**
(c) Both
(d) Neither

**16. Which of the following is an example of a disproportionation reaction?**

(a) 2H2O -> 2H2 + O2
✓ **(b) Cl2 + H2O -> HCl + HClO**
(c) Fe + CuSO4 -> FeSO4 + Cu
(d) Zn + HCl -> ZnCl2 + H2

**17. Which species is reduced in the following reaction? Zn + Cu2+ -> Zn2+ + Cu**

(a) Zn
✓ **(b) Cu2+**
(c) Cu

(d) $Zn^{2+}$

**18. Oxidation number of oxygen in $H_2O_2$ is:**

✓ **(a) -1**

(b) -2

(c) 0

(d) +1

**19. The oxidation number of sulfur in $H_2SO_4$ is:**

(a) +2

(b) +4

✓ **(c) +6**

(d) +8

**20. Redox reactions always involve:**

(a) Gain of protons

✓ **(b) Transfer of electrons**

(c) Loss of neutrons

(d) Nuclear change

**21. The reducing agent in a redox reaction:**

(a) Gains electrons

✓ **(b) Loses electrons**

(c) Accepts protons

(d) Donates protons

**22. In redox reactions, oxidizing agents are:**

✓ **(a) Reduced**

(b) Oxidized

(c) Always metals

(d) Always gases

**23. In a galvanic cell, the anode is:**

✓ **(a) Negative and site of oxidation**

(b) Positive and site of oxidation

(c) Negative and site of reduction

(d) Positive and site of reduction

**24. Electrons in a galvanic cell flow from:**

✓ **(a) Anode to cathode**

(b) Cathode to anode

(c) Salt bridge to cathode

(d) Electrolyte to anode

**25. Which component maintains electrical neutrality in an electrochemical cell?**

(a) Voltmeter

✓ **(b) Salt bridge**

(c) Electrolyte

(d) Electrodes

**26. Standard hydrogen electrode is assigned a potential of:**

(a) 1.0 V

(b) 0.5 V

✓ **(c) 0.0 V**

(d) -1.0 V

**27. The cathode in an electrolytic cell is:**

(a) Positive electrode
✓ **(b) Negative electrode**
(c) Neutral electrode
(d) None

**28. The anode in electrolysis of molten NaCl is:**

(a) Na+
✓ **(b) Cl-**
(c) Na
(d) H+

**29. In electrolysis, reduction always occurs at:**

(a) Anode
✓ **(b) Cathode**
(c) Salt bridge
(d) Electrolyte

**30. A Daniell cell produces electricity through:**

(a) Electrolysis
(b) Radioactivity
✓ **(c) Spontaneous redox reaction**
(d) Endothermic reaction

**31. A salt bridge is usually made from:**

(a) Metallic wires
(b) Porous glass
✓ **(c) Inert salt solution in agar-agar**
(d) Hydrochloric acid

**32. Cell potential is the difference between:**

(a) Temperature and pressure
(b) Concentrations of ions
✓ **(c) Electrode potentials of cathode and anode**
(d) Mass of electrodes

**33. Standard electrode potential is measured under:**

(a) 100 deg C and 1 atm
(b) 0 deg C and 1 M concentration
✓ **(c) 25 deg C and 1 M concentration**
(d) 100 deg C and 1 M concentration

**34. A more negative E deg value means the electrode:**

(a) Is a strong oxidizing agent
✓ **(b) Is a strong reducing agent**
(c) Is neutral
(d) Has high electronegativity

**35. Which metal is the best reducing agent?**

(a) Cu
(b) Zn
(c) Fe
✓ **(d) K**

**36. Which element has E deg = 0.00 V?**

(a) H+
(b) H2
(c) SHE
✓ **(d) All of the above**

**37. A spontaneous redox reaction occurs if:**
(a) E deg cell < 0
(b) E deg cell = 0
✓ **(c) E deg cell > 0**
(d) E deg cell is negative

**38. The E deg cell of a cell is calculated as:**
(a) E deg anode + E deg cathode
✓ **(b) E deg cathode - E deg anode**
(c) E deg anode - E deg cathode
(d) (E deg anode + E deg cathode) / 2

**39. Which of the following elements cannot displace hydrogen from acid?**
(a) Zn
(b) Al
✓ **(c) Cu**
(d) Fe

**40. Which metal lies below hydrogen in the electrochemical series?**
(a) Na
(b) Mg
(c) Zn
✓ **(d) Ag**

**41. Which of the following will react with HCl to liberate H2 gas?**
(a) Ag
(b) Au
✓ **(c) Mg**
(d) Cu

**42. The more positive the standard reduction potential:**
(a) The stronger the reducing agent
(b) The weaker the oxidizing agent
✓ **(c) The stronger the oxidizing agent**
(d) No effect

**43. Electrolysis is the process of:**
(a) Chemical energy into electrical
✓ **(b) Electrical energy into chemical**
(c) Heat energy into mechanical
(d) Mechanical energy into chemical

**44. Electrolysis requires:**
(a) Redox reaction
(b) Spontaneous cell
✓ **(c) Non-spontaneous reaction**
(d) Combustion

**45. Faraday's First Law states that mass deposited is directly proportional to:**
(a) Time

(b) Voltage
(c) Resistance
✓ **(d) Quantity of electricity passed**

**46. One Faraday is equal to:**
(a) 1 C
✓ **(b) 96,500 C**
(c) 1 J
(d) 96,500 J

**47. Electrochemical equivalent is:**
(a) Mass per mole
✓ **(b) Mass per coulomb**
(c) Charge per second
(d) Current per mass

**48. If 1 Faraday of electricity is passed, the mass deposited equals:**
✓ **(a) 1 gram equivalent**
(b) 1 gram
(c) 1 mole
(d) 1 atom

**49. Electroplating involves:**
(a) Spontaneous oxidation
(b) Redox without electrodes
✓ **(c) Electrolysis for coating**
(d) Thermolysis

**50. Electrolysis of CuSO4 using copper electrodes results in:**
(a) Increase in electrolyte concentration
(b) Decrease in Cu2+ concentration
✓ **(c) No change in electrolyte composition**
(d) Formation of new compound

**51. Electrolyte in electroplating of silver is:**
(a) NaCl
✓ **(b) AgNO3**
(c) CuSO4
(d) HNO3

**52. The number of coulombs required to deposit 1 mole of Ag+:**
✓ **(a) 96,500 C**
(b) 193,000 C
(c) 241,250 C
(d) 1 C

**53. The process of coating a metal with zinc is called:**
(a) Electroplating
(b) Electrorefining
✓ **(c) Galvanization**
(d) Alloying

**54. In electroplating, the object to be plated is connected to:**
✓ **(a) Cathode**
(b) Anode

(c) Both electrodes
(d) Electrolyte

**55. Electrolysis is used in the extraction of:**
(a) Gold
(b) Silver
✓ **(c) Aluminum**
(d) Mercury

**56. The process of purifying copper using electricity is called:**
(a) Electrolysis
(b) Electroplating
✓ **(c) Electrorefining**
(d) Electrogravimetry

**57. The metal deposited at cathode during electrolysis of NaCl is:**
(a) Cl2
(b) H2
✓ **(c) Na**
(d) Cu

**58. Which of the following is not a use of electrochemistry?**
(a) Electroplating
(b) Extraction of metals
✓ **(c) Manufacture of explosives**
(d) Electrorefining

**59. Electrolysis of brine produces:**
(a) Hydrogen, oxygen, NaOH
(b) Sodium, chlorine, H2O
✓ **(c) Hydrogen, chlorine, NaOH**
(d) NaOH, Cl2, CO2

**60. Downs cell is used for extraction of:**
(a) Zinc
(b) Iron
✓ **(c) Sodium**
(d) Aluminum

**61. In the Hall-Heroult process, aluminum is obtained by:**
(a) Chemical reduction
(b) Thermal decomposition
✓ **(c) Electrolysis of alumina**
(d) Roasting and leaching

**62. Electrochemical cells convert chemical energy into:**
(a) Heat
✓ **(b) Electrical energy**
(c) Light
(d) Nuclear energy

**63. If E deg cell is +1.10 V, the reaction is:**
✓ **(a) Spontaneous**
(b) Non-spontaneous
(c) At equilibrium

(d) Endothermic only

**64. What is the unit of cell potential?**

(a) Ampere
(b) Ohm
✓ **(c) Volt**
(d) Farad

**65. Which metal will displace Cu from CuSO4 solution?**

(a) Ag
✓ **(b) Zn**
(c) Hg
(d) Pb

**66. For a redox reaction to be spontaneous, Delta G should be:**

(a) Zero
(b) Positive
✓ **(c) Negative**
(d) Constant

**67. Which half-cell has the highest tendency to gain electrons?**

(a) Li+/Li
✓ **(b) F2/F-**
(c) Na+/Na
(d) Al3+/Al

**68. In an electrochemical cell, which ion migrates to cathode through the salt bridge?**

✓ **(a) Cations**
(b) Anions
(c) Electrons
(d) Protons

**69. Electrochemical series helps in predicting:**

(a) Rate of reaction
(b) Type of bond
✓ **(c) Direction of redox reactions**
(d) Heat released

**70. If E deg for Zn2+/Zn is -0.76 V and for Cu2+/Cu is +0.34 V, then E deg cell is:**

✓ **(a) +1.10 V**
(b) -1.10 V
(c) 0.42 V
(d) -0.42 V

**71. The reaction at the cathode is always:**

(a) Oxidation
(b) Neutral
✓ **(c) Reduction**
(d) Ionization

**72. An increase in temperature generally:**

(a) Increases cell voltage
✓ **(b) Decreases cell voltage**
(c) No effect
(d) Depends on pressure

# Section 2 | Short Questions

**Q6. Calculate the oxidation number of chromium (Cr) in: (i) CrCl3, (ii) Cr2(SO4)3, (iii) Cr2O7(2-)**

(i) CrCl3: x + 3(-1) = 0, so x = +3.

(ii) Cr2(SO4)3: sulfate charge -2 x 3 = -6; 2x + (-6) = 0, so x = +3.

(iii) Cr2O7(2-): oxygen charge -2 x 7 = -14; 2x + (-14) = -2, so 2x = +12, x = +6.

**Q7. Given reactivity order K > Mg > Zn > Fe > Cu, predict the reactions (if any) for: (i) Cu added to MgSO4 solution, (ii) Fe added to dilute HCl.**

(i) No reaction: copper is less reactive than magnesium and cannot displace it from solution.

(ii) Reaction occurs: Fe(s) + 2HCl(aq) -> FeCl2(aq) + H2(g), since iron is more reactive than hydrogen.

**Q8. Explain why some metals higher in the activity series can displace hydrogen from acids while others lower in the series cannot.**

Metals higher in the activity series are more reactive and lose electrons (oxidize) more readily; they displace H+ ions from acids because they have a stronger tendency to oxidize than hydrogen.

Metals lower in the series (e.g. Cu, Pt) have a weaker tendency to lose electrons and cannot displace hydrogen ions from acids.

**Q9. Calculate the number of Faradays required to deposit 108 g Ag+, 63.5 g Cu2+, and 27 g Al3+.**

Faradays = Mass / (Molar Mass x n)

Ag+ (n=1): 108/(108x1) = 1 Faraday

Cu2+ (n=2): 63.5/(63.5x2) = 0.5 Faraday

Al3+ (n=3): 27/(27x3) = 0.333 Faraday

Total = 1 + 0.5 + 0.333 = 1.833 Faradays

**Q10. A current of 0.500 A is passed through AgNO3 solution for 30.0 min; 0.503 g Ag is deposited. Molar mass of Ag = 107.87 g/mol. Calculate Avogadro's number.**

Q = I x t = 0.500 x 1800 = 900 C

Moles of Ag = 0.503/107.87 = 4.664x10^-3 mol; moles of electrons = moles of Ag (since Ag+ needs 1 electron)

F = Q/moles = 900/4.664x10^-3 = 1.93x10^5 C/mol

Na = F/e = (1.93x10^5)/(1.602x10^-19) = 6.02x10^23

**Q11. A cell has standard nickel (E deg = -0.25 V) and standard cobalt (E deg = -0.28 V) electrodes. Identify anode/cathode, write the cell reaction, and find E(cell) deg.**

Co has the more negative E deg, so it is oxidized and acts as the anode; Ni is the cathode.

Anode: Co(s) -> Co2+(aq) + 2e-; Cathode: Ni2+(aq) + 2e- -> Ni(s)

Overall: Co(s) + Ni2+(aq) -> Co2+(aq) + Ni(s)

E(cell) deg = E(cathode) - E(anode) = (-0.25) - (-0.28) = +0.53 V

**Q12. What is electrochemistry?**

Electrochemistry is the branch of chemistry dealing with the relationship between electrical energy and chemical reactions, including redox reactions where electrons are transferred.

Example: electrolysis of water produces hydrogen and oxygen gases using electricity.

**Q13. What is a redox reaction?**

A redox reaction involves the transfer of electrons - one substance is oxidized (loses electrons) and another is reduced (gains electrons).

Example: Zn(s)+Cu2+(aq) -> Zn2+(aq)+Cu(s) (Zn is oxidized; Cu2+ is reduced).

**Q14. What is oxidation in terms of electron transfer?**

Oxidation is the loss of electrons by an atom or ion.

Example: Na(s) -> Na+(aq)+e- (sodium loses one electron).

**Q15. What is reduction in terms of electron transfer?**

Reduction is the gain of electrons by an atom or ion.

Example: Cl2(g)+2e- -> 2Cl-(aq) (chlorine gains electrons).

**Q16. Write the nature of photosynthesis and respiration reactions.**

Photosynthesis is a redox reaction that provides food for the entire planet; respiration is a redox reaction that keeps us alive. Both are redox reactions.

**Q17. What is an electrochemical cell?**

A device that converts chemical energy into electrical energy through redox reactions.

Example: the Daniell cell (Zn-Cu cell).

**Q18. What is the function of a salt bridge in a galvanic cell?**

A salt bridge maintains electrical neutrality by allowing the movement of ions between the two half-cells.

Example: a KNO3 salt bridge connects the Zn and Cu compartments in a Daniell cell.

**Q19. What are electrodes?**

Conductors (usually metals) where oxidation or reduction occurs: the anode is where oxidation occurs, and the cathode is where reduction occurs.

**Q20. What is a galvanic cell?**

A cell that generates electrical energy from a spontaneous redox reaction.

Example: Daniell cell, Zn(s)/Zn2+(aq) || Cu2+(aq)/Cu(s).

**Q21. Which electrode is negative in a galvanic cell, and why?**

The anode is negative because it loses electrons (oxidation).

**Q22. What flows in the external circuit of a galvanic cell?**

Electrons flow from the anode to the cathode, and from there can be used in the external circuit.

**Q23. Why is the Daniell cell important in electrochemistry?**

It is a classic example of a galvanic cell that demonstrates the conversion of chemical energy into electrical energy.

**Q24. What is an electrolytic cell?**

A device where electrical energy is used to drive a non-spontaneous chemical reaction.

Example: electrolysis of molten NaCl.

**Q25. Which electrode is positive in an electrolytic cell?**

The anode is positive because it attracts negative ions (oxidation still occurs there).

**Q26. What happens during electrolysis of molten NaCl?**

At the cathode: Na+ + e- -> Na(s). At the anode: 2Cl- -> Cl2(g) + 2e-

**Q27. What is electroplating?**

The process of depositing a metal layer on an object using an electrolytic cell.

Example: electroplating silver onto a spoon.

**Q28. What is electrode potential?**

The potential difference between a metal electrode and its ion solution.

**Q29. What is standard electrode potential (E deg)?**

The electrode potential measured under standard conditions: 1 M concentration, 1 atm pressure, 25 deg C.

**Q30. What is the standard hydrogen electrode (SHE)?**

A reference electrode assigned E deg = 0.00 V, used to measure other electrode potentials.

**Q31. Why is SHE used as a reference?**

Because it is stable, reversible, and easy to reproduce.

**Q32. What is the EMF of a cell?**

The electromotive force is the potential difference between the electrodes of a cell when no current is drawn: E(cell) deg = E(cathode) deg - E(anode) deg

**Q33. How is cell potential related to spontaneity?**

A positive EMF indicates a spontaneous reaction; a negative EMF indicates a non-spontaneous reaction.

**Q34. How do you write cell notation for a galvanic cell?**

Anode / Anode solution || Cathode solution / Cathode

Example: Zn(s)/Zn2+(aq) || Cu2+(aq)/Cu(s)

**Q35. What is the electrochemical series?**

A list of elements arranged in order of increasing standard electrode potentials.

# Section 3 | Descriptive & Numerical Problems

## Descriptive Questions

**Q1 (Descriptive). What is the hit-and-trial method for balancing redox equations?**

Balance atoms of all elements except nitrogen, oxygen and hydrogen first; then balance nitrogen atoms (if present); then oxygen atoms; and lastly hydrogen atoms.

## Numerical Problems

**Q7 (Numerical). Calculate the electrode potential for a Zn electrode in 0.010 mol dm-3 ZnSO4 at 298 K. E deg(Zn2+/Zn) = -0.76 V. (R=8.31 J/K/mol, F=96500 C/mol)**

Half-reaction: Zn2+(aq)+2e- -> Zn(s), n=2

Nernst equation: E = E deg - (RT/nF) x ln Q, where Q = 1/[Zn2+] = 1/0.010 = 100

(8.31x298)/(2x96500) = 0.012837 V; ln 100 = 4.606

E = -0.76 - (0.012837x4.606) = -0.76 - 0.0591 = -0.82 V

**Q8 (Numerical). A constant current of 2.00 A is passed through CuSO4 solution for 30.0 min. Calculate the mass of copper deposited. (M(Cu)=63.5 g/mol, F=96500 C/mol)**

Cathode: Cu2+ + 2e- -> Cu(s); Q = I x t = 2.00 x 1800 = 3600 C

Moles electrons = Q/F = 3600/96500 = 0.0373 mol; moles Cu = 0.0373/2 = 0.01865 mol

Mass = moles x molar mass = 0.01865 x 63.5 = 1.18 g of copper

**Q9 (Numerical). A galvanic cell has SHE and a Ni2+/Ni half-cell. Measured potential at 298 K is 0.25 V, with Ni as the negative terminal. Find (a) the cell reaction, (b) E deg(Ni2+/Ni), (c) anode and cathode.**

(a) SHE: 2H+(aq)+2e- -> H2(g); Ni is the anode (oxidation): Ni(s) -> Ni2+(aq)+2e-. Overall: Ni(s)+2H+(aq) -> Ni2+(aq)+H2(g)

(b) E(cell) deg = E(cathode) - E(anode); 0.25 = 0.00 - E(Ni2+/Ni), so E(Ni2+/Ni) = -0.25 V

(c) Anode: Ni(s) (negative terminal); Cathode: SHE (H2/H+, positive terminal)

# Section 4 | Long Questions

## Q1. What is electrochemistry? Explain oxidation, reduction and redox reactions.

### Electrochemistry

Definition: the branch of chemistry that deals with the interconversion of electrical and chemical energy.

### Oxidation, Reduction, and Redox Reactions

Oxidation: a process involving the loss of one or more electrons. Examples: Fe2+ -> Fe3+ + e-; Zn0 -> Zn2+ + 2e- (both lose electrons and are oxidized).

Reduction: a process involving the gain of one or more electrons. Examples: Cl0 + e- -> Cl-; Cu2+ + 2e- -> Cu0 (both gain electrons and are reduced).

Redox Reaction: a reaction in which oxidation and reduction both take place. Example: Zn0 + Cu2+ -> Cu0 + Zn2+ (zinc is oxidized from 0 to +2; copper is reduced from +2 to 0).

### Identifying Oxidation and Reduction

Two ways to determine whether a substance has been oxidized or reduced: (1) electron transfer, (2) changes in oxidation number.

Photosynthesis is a redox reaction that provides food for the entire planet; respiration is a redox reaction that keeps us alive.

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## Q2. Why is it important to assign oxidation numbers? Explain the rules for assigning oxidation number.

### Oxidation Number and Its Significance

Definition: the apparent charge on an atom of an element in a molecule or ion - also called oxidation state; it measures the gain or loss of electrons by an element.

Oxidation numbers can be positive, negative, or zero, and the sign must always be included. A higher positive number means more oxidized; a higher negative number means more reduced.

### Rules for Assigning Oxidation Numbers

(i) Uncombined elements: oxidation number is always zero (e.g. S8, H2, Cl2, Na, Mg, Zn in elemental form).

(ii) Electronegativity rule: the more electronegative element in a compound/ion is given a negative oxidation number (e.g. in HF, F is -1, H is +1).

(iii) Fixed oxidation numbers: Group 1 metals always +1; Group 2 metals always +2; Group 17 halogens -1 in binary compounds; hydrogen usually +1 (but -1 in metal hydrides like NaH); oxygen usually -2 (but -1 in peroxides like H2O2, and +2 in F2O).

(iv) Monoatomic ions: oxidation number equals the ion's charge (e.g. Cl- = -1, Al3+ = +3).

### Worked Examples

S in SO2: O=-2 each, 2 O gives -4; S+(-4)=0, so S=+4.

S in SO4(2-): O=-2 each, 4 O gives -8; S+(-8)=-2, so S=+6.

Mn in KMnO4: K=+1, O=-2 each (4 O = -8); +1+Mn-8=0, so Mn=+7.

## Q4. Define and explain disproportionation reaction. Explain oxidizing and reducing agents.

### Disproportionation Reaction

Definition: a reaction where a single substance acts as both the oxidizing and reducing agent, giving two different products with different oxidation states.

Example 1: decomposition of hydrogen peroxide - 2H2O2(l) -> 2H2O(l)+O2(g).

Example 2: 3Cl2+6NaOH -> 5NaCl+NaClO3+3H2O. Cl in Cl2 = 0; Cl in NaCl = -1 (reduced); Cl in NaClO3 = +5 (oxidized) - a classic disproportionation of chlorine.

### Oxidizing and Reducing Agents

Oxidizing Agent (oxidant): a substance that oxidizes another substance and is itself reduced by gaining electrons. General form: X(0) + e- -> X-(lower). Example: Cl + e- -> Cl- (Cl acts as oxidizing agent).

Reducing Agent (reductant): a substance that reduces another substance and is itself oxidized by losing electrons. General form: M(0) -> M+ + e- (higher). Example: Na -> Na+ + e- (Na acts as reducing agent).

### Quick Check 10.2 - Identifying Species

Fe2O3(s)+CO(g) -> Fe(s)+CO2(g): CO is oxidized (C: +2 to +4) and is the reducing agent; Fe2O3 is reduced (Fe: +3 to 0) and is the oxidizing agent.

CH4(g)+2O2(g) -> CO2(g)+2H2O(l): CH4 is oxidized (C: -4 to +4) and is the reducing agent; O2 is reduced (0 to -2) and is the oxidizing agent.

In 2Na2O2+2H2O -> 4NaOH+O2: oxygen in Na2O2 (-1) is both reduced (to -2 in NaOH) and oxidized (to 0 in O2) - a disproportionation.

## Q5. Explain the rules for balancing the redox equation by the oxidation number method.

### Rules for Balancing by Oxidation Number Method

(i) Write the skeleton equation of the reaction.

(ii) Identify elements that undergo a change in oxidation number.

(iii) Record the oxidation number above the symbols of those elements.

(iv) Indicate the change with arrows joining atoms on both sides, showing electrons gained or lost.

(v) Equate the increase and decrease in oxidation number (electrons gained = electrons lost) by multiplying with suitable coefficients.

(vi) Balance the rest of the equation by inspection (hit and trial).

### Worked Balances

Fe2O3+3CO -> 2Fe+3CO2 (Fe: +3->0 gain 3e per Fe x2 = 6e; C: +2->+4 lose 2e per C x3 = 6e - balanced).

MnO2+4HCl -> MnCl2+2H2O+Cl2 (Mn: +4->+2 gain 2e; 2Cl: -1->0 lose 2e - balanced).

5Fe2+ + MnO4- + 8H+ -> 5Fe3+ + Mn2+ + 4H2O (Fe: +2->+3 lose 1e x5=5e; Mn: +7->+2 gain 5e - balanced).

K2Cr2O7+6FeSO4+7H2SO4 -> Cr2(SO4)3+3Fe2(SO4)3+K2SO4+7H2O (Cr: +6->+3 gain 3e x2=6e; Fe: +2->+3 lose 1e x6=6e - balanced).

## Q6. What is the construction and working of an electrolytic cell? Explain the redox reaction in it.

### Electrolytic Cell

Definition: a device that converts electrical energy into chemical energy through electrolysis.

### Construction and Working

At the cathode (-): reduction happens - cations move to the cathode and gain electrons; products may be metals or hydrogen gas. Examples: Cu2+(aq)+2e- -> Cu(s); 2H+(aq)+2e- -> H2(g).

At the anode (+): oxidation happens - anions move to the anode and lose electrons; products may be non-metal gases or other substances. Examples: 2Cl-(aq) -> Cl2(g)+2e-; 4OH-(aq) -> O2(g)+2H2O(l)+4e-.

### Example: Electrolysis of Molten ZnCl2

Overall reaction: ZnCl2(l) -> Zn(s)+Cl2(g). Electrons lost at the anode equal electrons gained at the cathode, maintaining charge balance.

### Quick Check 10.4 - Electrolysis of Molten CuBr2

(i) Ions present: Cu2+ and Br- (molten, not aqueous).

(ii) Copper metal forms at the cathode (Cu2+ gains electrons - reduction).

(iii) Cu2+ + 2e- -> Cu (reduction at cathode).

(iv) 2Br- -> Br2 + 2e- (oxidation at anode).

## Q7. Explain the mass of a substance deposited during electrolysis and the amount of substance formed.

### Mass Deposited During Electrolysis

The mass deposited (or liberated) is directly proportional to the time current flows, the strength of the current, and the number of electrons involved in the electrode reaction.

Total charge: Q = I x t, where Q is in coulombs, I in amperes, t in seconds. 1 Faraday (F) = 96,500 C/mol.

### Examples

Electrolysis of AgNO3: Ag+(aq)+e- -> Ag(s). 1 Faraday deposits 1 mole of silver (107.9 g); charge required = 1x96,500 C.

Electrolysis of CuSO4: Cu2+(aq)+2e- -> Cu(s). 2 Faradays deposit 1 mole of copper (63.5 g); charge required = 2x96,500 = 193,000 C.

General formula: No. of Faradays = No. of moles of electrons gained or lost.

### Sample Problem 10.5

Calculate the mass of Pb deposited when 1.50 A flows through molten PbBr2 for 20.0 min. (Ar[Pb]=207, F=96500 C/mol)

Half-equation: Pb2+(aq)+2e- -> Pb(s); 2 moles electrons per mole Pb = 2F = 193,000 C/mol.

Q = I x t = 1.50x1200 = 1800 C.

Mass = (207/193000)x1800 = 1.93 g of Pb

## Q8. How can Avogadro's number be derived using an electrolytic cell?

### Avogadro Constant and Electrolysis

The electrolytic method allows determination of Na by calculating the charge on 1 mole of electrons and dividing by the charge on a single electron.

Na = (Charge on 1 mole of electrons)/(Charge on 1 electron) = 96,500/(1.60x10^-19) = 6.03x10^23 mol-1

### Experimental Determination

Objective: calculate the charge carried by 1 mole of electrons using electrolysis of CuSO4 solution with copper electrodes.

Apparatus: pure copper cathode and anode, aqueous CuSO4, power supply, variable resistor, ammeter, stopwatch, distilled water and propanone for cleaning, electronic balance.

Procedure: weigh the electrodes before electrolysis; pass a constant current (e.g. 0.2 A) for a fixed time (e.g. 40 minutes = 2400 s); after electrolysis, clean, dry, and reweigh both electrodes to find the mass change.

### Quick Check 10.5

(a) 0.35 A passed through AgNO3 for 18 min: Q = 0.35x1080 = 378 C; Mass = (Q x Ar)/(n x F) = (378x108)/(1x96500) = 0.423 g Ag deposited.

(b) 1.04 A through dilute H2SO4 for 6.00 min gave 41.5 cm3 H2 at STP: Q = 1.04x360 = 374.4 C. To liberate 1 mole H2 (2H++2e- -> H2), charge required = 2F = 193,000 C.

## Q9. Explain the electrode potential for reactive and unreactive metals.

### Electrode Potential

Definition: when a metal is dipped in a solution of its own ions, an electrical potential is established between the metal and its ions - this is electrode potential, indicating the ease of oxidation or reduction.

Example: Cu2+(aq)+2e- <=> Cu(s). Opposing reactions occur at the electrode - oxidation (Cu -> Cu2++2e-) and reduction (Cu2++2e- -> Cu). Equilibrium is reached when the oxidation rate equals the reduction rate.

### Comparison of Metals

Unreactive metals (e.g. copper): the redox equilibrium lies to the right (Cu2+ ions are readily reduced), giving a higher electrode potential.

Reactive metals (e.g. vanadium): the redox equilibrium lies to the left (V2+ ions are harder to reduce), giving a lower electrode potential.

### Electrical Double Layer

Absolute electrode potentials cannot be measured directly because of the electrical double layer formed when a metal is dipped in a solution of its own ions.

Example: Zn(s) <=> Zn2+(aq)+2e-. Zn2+ ions accumulate near the metal surface and are attracted back, creating a potential difference that cannot be measured directly.

The Standard Hydrogen Electrode (SHE) is used as a reference/comparison system to measure electrode potential in volts.

## Q10. What are the standard conditions for measuring standard electrode potential? How is it measured?

### Standard Electrode Potential (E deg)

Definition: the voltage of a half-cell measured under standard conditions, with the standard hydrogen electrode (SHE) as the other half-cell. It reflects a substance's tendency to gain or lose electrons.

Purpose: compares oxidizing/reducing power of different electrodes. More positive E deg = stronger tendency to gain electrons (reduction); more negative E deg = stronger tendency to lose electrons (oxidation).

### Types of Half-Cells

(i) Metal/Metal ion half-cell: e.g. Cu2+(aq)+2e- -> Cu(s), E deg = +0.34 V.

(ii) Non-metal/Non-metal ion half-cell: e.g. Cl2(g)+2e- -> 2Cl-(aq), E deg = +1.36 V.

(iii) Ion/Ion half-cell: e.g. Fe3+(aq)+e- -> Fe2+(aq), E deg = +0.77 V.

### Measurement Examples

Cu2+/Cu connected to SHE via a salt bridge gives +0.34 V - Cu2+ ions are more likely to gain electrons than H+ ions.

Zn2+/Zn connected to SHE gives -0.76 V - Zn2+ has a lower (more negative) E deg, so zinc is oxidized, giving up electrons to the H+/H2 half-cell.

## Q12. Describe the construction and working principle of the electrochemical (galvanic) cell.

### Galvanic (Voltaic) Cell

Definition: a cell in which chemical energy converts spontaneously into electrical energy via an exothermic redox process, with oxidation at the anode and reduction at the cathode.

Cell representation: Cu2+/Cu || Zn2+/Zn cell, E deg = 1.10 V.

### Construction

Two half-cells connected by wires (to a high-resistance voltmeter) and a salt bridge that completes the circuit by allowing ion movement while preventing electron flow and charge buildup - often made from filter paper soaked in saturated KNO3 solution.

### Working

At the anode: Zn(s) -> Zn2+(aq)+2e- (oxidation). At the cathode: Cu2+(aq)+2e- -> Cu(s) (reduction).

Overall: Zn(s)+Cu2+(aq) -> Zn2+(aq)+Cu(s); E(cell) deg = E(red)+E(oxi) = 0.34-(-0.76) = +1.10 V.

In this cell, zinc is the anode and the Cu2+/Cu couple is the cathode.

### Quick Check 10.6

(a) The salt bridge completes the circuit by allowing ion flow, maintains electrical neutrality, prevents direct mixing of solutions, and keeps the cell running continuously.

(b) For Fe2+/Fe (E deg=-0.44V) and Cu2+/Cu (E deg=+0.34V): Fe is the anode, Cu is the cathode; E(cell) deg = 0.34-(-0.44) = 0.78 V.

## Q13. Explain the applications of standard electrode potential values.

### (i) Direction of Electron Flow

Comparing E deg values of two half-cells shows the direction of electron flow. Example: Ag+/Ag (E deg=+0.80V) and Zn2+/Zn (E deg=-0.76V) - electrons flow from Zn (more negative, loses electrons more readily) to Ag+ (more positive, gains electrons more readily). Rule: electrons flow from the anode to the cathode, i.e. from more negative to more positive E deg.

### (ii) Feasibility of a Reaction

More positive E deg = greater tendency for the forward (reduction) direction = stronger oxidizing agent; more negative E deg = greater tendency for reverse (oxidation) = stronger reducing agent.

Example 1: Mg2+ + Ag -> ? (Mg2+/Mg E deg=-2.32V, Ag+/Ag E deg=+0.80V): E(cell)=(-2.32)-(0.80)=-1.52V -> not feasible.

Example 2: Ag+ + Mg -> ?: E(cell)=0.80-(-2.32)=+1.52V -> feasible: 2Ag+(aq)+Mg(s) -> Mg2+(aq)+2Ag(s)

## Q15. Write a detailed note on the Nernst equation to explain the variation of electrode potential.

### Nernst Equation

The effect of concentration and temperature on E(cell) can be deduced using the Nernst equation.

For an electrode: E = E deg + (RT/zF) x ln([Oxidized form]/[Reduced form]), where R=8.314 J/K/mol, T=temperature (K), z=electrons transferred, F=Faraday constant.

Simplified form at 25 deg C: E = E deg + (0.059/z) x log([Oxidized form]/[Reduced form])

### Sample Problem 10.6

Electrode potential of Zn in 0.1 M Zn2+ solution, E deg(Zn2+/Zn) = -0.76 V.

E = (-0.76) + [(8.31x298)/(2x96500)] x ln(0.1) = -0.76 + (-0.029) = -0.78 V

Conclusion: as Zn2+ concentration decreases, the equilibrium shifts backward and electrode potential decreases (becomes more negative).

## Q16. How does ease of oxidation relate to standard electrode potential? Discuss the activity series of metals.

### Activity Series

The activity series ranks metals by reactivity, particularly their tendency to lose electrons (oxidize).

Highly reactive metals: have more negative standard reduction potentials, lose electrons readily, and act as reducing agents - e.g. alkali metals like cesium and sodium. Example: 2Na+2H2O -> 2NaOH+H2 (up arrow).

Least reactive metals: do not oxidize easily - noble metals like gold and silver are least active.

### Uses of the Activity Series

Predicts metal displacement reactions, reactivity with water and acids, and whether a redox reaction will occur - a practical tool in electrochemistry, metallurgy, and industrial metal extraction.

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## Q17. What does it mean for a redox reaction to be feasible or spontaneous?

### Feasibility from Activity Series / E deg Data

A more reactive metal (higher in the activity series) can reduce the ion of a less reactive metal. If E(cell) deg is positive, the reaction is feasible; if negative, it is not.

Steps: write half-reactions; look up E deg values; calculate E(cell) deg; interpret - E(cell)>0 feasible, E(cell)<0 not feasible.

### Worked Examples

Fe2+/Fe (E deg=-0.44V) and Ag+/Ag (E deg=+0.80V): overall Fe(s)+2Ag+(aq) -> Fe2+(aq)+2Ag(s); E(cell)=0.80-(-0.44)=+1.24V -> feasible.

Al3+/Al (E deg=-1.66V) and Ag+/Ag (E deg=+0.80V): overall Al3+(aq)+Ag(s) -> Al(s)+Ag+(aq); E(cell)=-1.66+0.80=-0.86V -> not feasible.

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## Q18. How does a photovoltaic cell convert light energy into electrical energy?

### Photovoltaic Cell

Definition: a device that converts light energy directly into electrical energy using the photovoltaic effect - the generation of voltage and current in a material exposed to light.

### Working

1. Light absorption: sunlight (photons) striking a semiconductor transfers energy to electrons.

2. Excitation: photon energy excites electrons from the valence band to the conduction band, creating free electrons and holes.

3. PN Junction: separates electrons and holes, preventing recombination.

4. Electric field: at the PN junction drives electrons to the n-side and holes to the p-side, creating a potential difference (voltage).

5. Current flow: connecting an external circuit lets electrons flow, producing electric current.

### Merits

Continues generating electricity as long as light is present - a sustainable, renewable energy source independent of fossil fuels.

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## Q19. Explain the principle behind the Winkler method for measuring dissolved oxygen.

### Biochemical Oxygen Demand (BOD)

Definition: the amount of oxygen consumed by microorganisms to decompose organic matter in water over 5 days at 25 deg C.

Significance: reflects organic pollution level in water - higher BOD means more pollution. Measured from the difference in dissolved oxygen (DO) before and after 5 days.

### Winkler Method for DO Measurement

Principle: dissolved oxygen oxidizes potassium iodide (KI) to release iodine.

Process: add MnSO4 and alkaline KI to the sample; oxygen reacts to form a brown MnO(OH)2 precipitate; acidify to release iodine; titrate the iodine with sodium thiosulfate using starch indicator (blue colour indicates iodine present, and its disappearance marks the endpoint).

More iodine formed means higher DO. There is an inverse relationship between BOD and DO: higher BOD means lower DO, and vice versa.

# Section 5 | Answer Key & Formula Summary

## MCQ Answer Key (Q1-Q72)

| Q.No | Ans | Q.No | Ans | Q.No | Ans | Q.No | Ans |
|---|---|---|---|---|---|---|---|
| 1 | C | 19 | C | 37 | C | 55 | C |
| 2 | C | 20 | B | 38 | B | 56 | C |
| 3 | B | 21 | B | 39 | C | 57 | C |
| 4 | B | 22 | A | 40 | D | 58 | C |
| 5 | C | 23 | A | 41 | C | 59 | C |
| 6 | C | 24 | A | 42 | C | 60 | C |
| 7 | D | 25 | B | 43 | B | 61 | C |
| 8 | B | 26 | C | 44 | C | 62 | B |
| 9 | B | 27 | B | 45 | D | 63 | A |
| 10 | B | 28 | B | 46 | B | 64 | C |
| 11 | C | 29 | B | 47 | B | 65 | B |
| 12 | D | 30 | C | 48 | A | 66 | C |
| 13 | C | 31 | C | 49 | C | 67 | B |
| 14 | B | 32 | C | 50 | C | 68 | A |
| 15 | B | 33 | C | 51 | B | 69 | C |
| 16 | B | 34 | B | 52 | A | 70 | A |
| 17 | B | 35 | D | 53 | C | 71 | C |
| 18 | A | 36 | D | 54 | A | 72 | B |

## Important Formulas

| Formula | Description |
|---|---|
| **Q = I x t** | Charge passed (Coulombs) |
| **1 Faraday (F) = 96,500 C/mol** | Charge on one mole of electrons |
| **No. of Faradays = Mass / (Molar Mass x n)** | Faradays required to deposit a substance |
| **Mass = (Q x Ar) / (n x F)** | Mass deposited during electrolysis |
| **Na = F / e** | Avogadro's Number from Faraday's constant |
| **E(cell) = E(cathode) - E(anode)** | Standard Cell Potential |
| **E = E deg + (RT/nF) x ln([Oxidized]/[Reduced])** | Nernst Equation |
| **E = E deg + (0.059/n) x log([Oxidized]/[Reduced])** | Nernst Equation at 25 deg C |
| **Delta G = -nFE(cell)** | Gibbs Free Energy and Cell Potential |

## Key Constants & Notes

| Constant/Note | Value |
|---|---|
| **Faraday's Constant (F)** | 96,500 C/mol |
| **Gas Constant (R)** | 8.31 J/K/mol |
| **Charge on Electron (e)** | 1.602 x 10^-19 C |
| **Avogadro's Number (Na)** | 6.02 x 10^23 mol-1 |
| **Standard Temperature** | 25 deg C / 298 K |
| **Standard Pressure** | 1 atm / 101 kPa |
| **Standard Hydrogen Electrode (SHE)** | E deg = 0.00 V |

## Formula Quick-Reference (Application Table)

| Formula | Application |
|---|---|
| **Oxidation Number Calculation** | Sum of oxidation numbers = 0 (neutral compound) or = charge (ion) |
| **Faradays Required** | Faradays = Mass / (Molar Mass x n) |
| **Charge (Q)** | Q = I x t |
| **Moles of Electrons** | Moles of e- = Q / F |
| **Mass Deposited** | Mass = (Q x M) / (n x F) |
| **Nernst Equation** | E = E deg - (RT/nF) x ln Q |
| **Cell Potential** | E(cell) deg = E(cathode) deg - E(anode) deg |
| **Avogadro's Number** | Na = F / e |

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