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Electrochemistry — CBSE Class 12 Chemistry Chapter Test

FREE CHAPTER TEST · CLASS XII · CHEMISTRY (043) · PATTERN OF THE CBSE SAMPLE QUESTION PAPER 2026-27

TIME ALLOWED	MAXIMUM MARKS	QUESTIONS	NAME / ROLL NO.
1½ hours	37	17 (A 8 · B 3 · C 3 · D 1 · E 2)	

GENERAL INSTRUCTIONS

Read the following instructions very carefully and strictly follow them:

1. This question paper contains 17 questions. All questions are compulsory.
2. This question paper is divided into five sections: A, B, C, D and E.
3. Section A (Q1–8) contains 6 multiple choice questions and 2 assertion-reason based questions of 1 mark each.
4. Section B (Q9–11) contains short answer questions of 2 marks each.
5. Section C (Q12–14) contains short answer questions of 3 marks each.
6. Section D (Q15) contains case-based questions of 4 marks each, with sub-parts of 1, 1 and 2 marks.
7. Section E (Q16–17) contains long answer questions of 5 marks each.
8. There is no overall choice. However, an internal choice has been provided in 1 question in Section B, 1 question in Section C, 1 question in Section D, 2 questions in Section E. Attempt only one of the alternatives in such questions.
9. Use of log tables and calculators is not allowed. You may use: F = 96500 C mol^{-1}, 2.303RT/F = 0.059 V at 298 K.
10. This chapter test follows the section structure, question types and marks per question of the CBSE Class XII Chemistry (043) Sample Question Paper 2026-27, at about half the length of the 70-mark board paper: 37 marks in 1½ hours.
11. The answer key and the marking scheme are at the end of this booklet. Attempt the paper in 1½ hours before you look at them.

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Chapter: Electrochemistry · **Topics:** Galvanic cells and electrode potentials · Nernst equation, ΔG° and K · Conductance and molar conductivity · Kohlrausch's law · Electrolysis and Faraday's laws · Batteries, fuel cells and corrosion

SECTION A

This section comprises 8 questions of 1 mark each: 6 multiple choice questions (MCQs) and 2 assertion-reason based questions.

1. The standard electrode potentials of Zn^{2+}/Zn and Cu^{2+}/Cu are −0.76 V and +0.34 V. The standard emf of the Daniell cell is 1

(A) 0.42 V (B) 1.10 V

(C) −1.10 V (D) 0.76 V

2. The SI unit of molar conductivity is 1

(A) S m^{-1} (B) S cm^{-1}

(C) Ω m (D) S m^2 mol^{-1}

3. The charge needed to deposit 1 mol of aluminium from Al^{3+} ions is 1

(A) 3 F (B) 1 F

(C) 2 F (D) 6 F

4. On dilution, the largest increase in molar conductivity is shown by an aqueous solution of 1

(A) KCl (B) NaCl

(C) CH_3COOH (D) HCl

5. In the rusting of iron, the reaction at the anodic spot is 1

(A) $Fe \rightarrow Fe^{2+} + 2e^-$ (B) $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

(C) $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ (D) $Fe^{3+} + e^- \rightarrow Fe^{2+}$

6. When a lead storage battery is discharged, the substance formed at both electrodes is 1

(A) Pb (B) PbO_2

(C) H_2SO_4 (D) $PbSO_4$

7. **Assertion (A) :** The molar conductivity of a weak electrolyte increases sharply on dilution. 1
Reason (R) : On dilution, the number of ions per unit volume of the solution increases.

(A) Both A and R are true, and R is the correct explanation of A.

(B) Both A and R are true, and R is not the correct explanation of A.

(C) A is true but R is false.

(D) A is false but R is true.

8. **Assertion (A) :** In a fuel cell, the chemical energy of the fuel is converted directly into electrical energy. 1
Reason (R) : A mercury cell gives a nearly constant voltage throughout its life.

(A) Both A and R are true, and R is the correct explanation of A.

(B) Both A and R are true, and R is not the correct explanation of A.

(C) A is true but R is false.

(D) A is false but R is true.

SECTION B

This section comprises 3 short answer questions of 2 marks each.

9. Find the standard Gibbs energy change for the cell reaction $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$, for which $E^\circ_{cell} = 0.46$ V. ($F = 96500$ C mol^{-1}) 2

10. (a) State Kohlrausch's law of independent migration of ions. Use it to find Λ°_m of CH_3COOH from the following values of Λ°_m (S cm^2 mol^{-1}): HNO_3 421, CH_3COOK 115, KNO_3 145. 2

OR

(b) Give two methods of preventing the corrosion of iron. Explain how galvanisation protects iron even when the coating is scratched.

11. A conductivity cell filled with a KCl solution of conductivity 0.012 S cm^{-1} has a resistance of 100 Ω. The same cell filled with another solution has a resistance of 400 Ω. Find the cell constant and the conductivity of the second solution. 2

SECTION C

This section comprises 3 short answer questions of 3 marks each.

12. (a) Find the emf of the cell $Zn(s) \mid Zn^{2+}(0.01\ M) \parallel Cu^{2+}(0.1\ M) \mid Cu(s)$ at 298 K. ($E^\circ_{cell} = 1.10$ V; take $2.303RT/F = 0.059$ V) 3

OR

(b) The standard emf of a cell in which two electrons are transferred is 0.59 V at 298 K. Find the equilibrium constant of the cell reaction and its standard Gibbs energy change. ($2.303RT/F = 0.059$ V; $F = 96500$ C mol^{-1})

13. A current of 2.0 A is passed through a solution of $CuSO_4$ for 965 s. Find (i) the charge passed, (ii) the mass of copper deposited at the cathode and (iii) the time needed to deposit 1.27 g of copper with the same current. (Cu = 63.5 g mol^{-1}; $F = 96500$ C mol^{-1}) 3

14. The conductivity of 0.002 M ethanoic acid is 6.0×10^{-5} S cm^{-1}. Its limiting molar conductivity is 400 S cm^2 mol^{-1}. Find its molar conductivity, degree of dissociation and dissociation constant. 3

SECTION D

This section comprises 1 case-based question of 4 marks, with sub-parts of 1, 1 and 2 marks.

15. Sodium hydroxide and chlorine are made on a large scale by the electrolysis of concentrated aqueous sodium chloride (brine). Although water could be oxidised at the anode, in practice chlorine is given off there, because the oxidation of water needs a large overpotential. Hydrogen is given off at the cathode, and sodium hydroxide is left in the solution.

(I) Name the gas given off at the anode when brine is electrolysed. [1]

(II) Why does the solution become alkaline as the electrolysis proceeds? [1]

(III) Write the reactions at the cathode and the anode in the electrolysis of brine, and explain why sodium metal is not deposited at the cathode. [2]

OR

(III) Name the products at the cathode and the anode when molten sodium chloride is electrolysed, and write the reactions at the two electrodes. [2]

SECTION E

This section comprises 2 long answer questions of 5 marks each.

16. (a) (i) Describe the Daniell cell, writing the reactions at the two electrodes. Give two functions of the salt bridge. 5

(ii) Write the Nernst equation for the cell Ni(s) | Ni^{2+}(aq) || Ag^{+}(aq) | Ag(s).

(iii) Find its emf at 298 K when [Ni^{2+}] = 0.1 M and [Ag^{+}] = 0.01 M. ($E°$: Ni^{2+}/Ni = −0.25 V, Ag^{+}/Ag = +0.80 V; 2.303RT/F = 0.059 V)

OR

(b) (i) Define conductivity and molar conductivity. How do they change on dilution? Explain the different ways in which the molar conductivity of strong and weak electrolytes changes with concentration.

(ii) A 0.05 M NaCl solution has a conductivity of 5.5×10^{-3} S cm^{-1}. Find its molar conductivity, and its resistance in a cell of cell constant 1.1 cm^{-1}.

17. (a) (i) Distinguish between primary and secondary batteries, giving one example of each. Write the reactions at the anode and the cathode of a lead storage battery while it is being used. 5

(ii) Write the electrode reactions in a hydrogen-oxygen fuel cell, and give two advantages of fuel cells.

OR

(b) (i) State Faraday's two laws of electrolysis.

(ii) Two cells containing $AgNO_3$ and $CuSO_4$ solutions are connected in series, and a current of 0.50 A is passed until 1.08 g of silver is deposited. Find the charge passed, the mass of copper deposited and the time for which the current flowed. (Ag = 108, Cu = 63.5 g mol^{-1}; F = 96500 C mol^{-1})

END OF THE QUESTION PAPER

ANSWER KEY — SECTION A

Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8
(B)	(D)	(A)	(C)	(A)	(D)	(C)	(B)

MARKING SCHEME AND WORKED SOLUTIONS

Marks for each step are shown in brackets, as in the CBSE marking scheme. Any correct alternative method earns full marks.

Q1. Answer (B). $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0.34 - (-0.76) = 1.10$ V.

Q2. Answer (D). $\Lambda_m = \kappa/c$: $(S\ m^{-1})/(mol\ m^{-3}) = S\ m^2\ mol^{-1}$.

Q3. Answer (A). $Al^{3+} + 3e^- \rightarrow Al$: 3 mol of electrons, i.e. 3 F.

Q4. Answer (C). Ethanoic acid is a weak electrolyte: its degree of dissociation rises sharply on dilution.

Q5. Answer (A). Iron is oxidised at the anodic spot; oxygen is reduced at the cathodic spot.

Q6. Answer (D). Anode: $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$; cathode: $PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$.

Q7. Answer (C). A is true (the degree of dissociation increases a lot). R is false: the number of ions per unit volume decreases on dilution. Answer C.

Q8. Answer (B). Both are true, but the steady voltage of a mercury cell (no ions in its overall reaction) has nothing to do with how a fuel cell works. Answer B.

Q9. (2 marks) Answer: -88.8 kJ mol^{-1}.
$\Delta_r G^\circ = -nFE^\circ_{cell}$ with n = 2 (each Cu atom gives two electrons to two Ag^+ ions). [1]
$\Delta_r G^\circ = -2 \times 96500 \times 0.46 = -88\,780$ J mol^{-1} ≈ -88.8 kJ mol^{-1}. [1]

Q10. (2 marks) Answer: (a) 391 S cm^2 mol^{-1}; (b) zinc is oxidised in preference to iron.
(a) At infinite dilution, the limiting molar conductivity of an electrolyte is the sum of the limiting molar conductivities of its cation and anion. [1]
$\Lambda^\circ(CH_3COOH) = \Lambda^\circ(HNO_3) + \Lambda^\circ(CH_3COOK) - \Lambda^\circ(KNO_3) = 421 + 115 - 145 = 391$ S cm^2 mol^{-1}. [1]
OR (b) Any two, e.g. painting or greasing (keeps out air and moisture); electroplating with a less reactive metal such as tin. [1] Galvanisation: coating iron with zinc. Zinc is more reactive (more negative E°), so it is oxidised in preference to iron and protects it even where the coating is scratched (sacrificial protection). [1]

Q11. (2 marks) Answer: 1.2 cm^{-1}; 3.0×10^{-3} S cm^{-1}.
Cell constant $G^* = \kappa R = 0.012 \times 100 = 1.2$ cm^{-1}. [1]
$\kappa = G^*/R = 1.2/400 = 3.0 \times 10^{-3}$ S cm^{-1}. [1]

Q12. (3 marks) Answer: (a) ≈ 1.13 V; (b) 10^{20}; ≈ -113.9 kJ mol^{-1}.
(a) Cell reaction: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$, n = 2. [1]
$E_{cell} = E^\circ_{cell} - (0.059/2)\log([Zn^{2+}]/[Cu^{2+}]) = 1.10 - 0.0295 \log(0.01/0.1)$ [1] $= 1.10 + 0.0295 \approx 1.13$ V. [1]
OR (b) $\log K_c = nE^\circ/0.059 = 2 \times 0.59/0.059 = 20$, so $K_c = 10^{20}$. [2]
$\Delta_r G^\circ = -nFE^\circ = -2 \times 96500 \times 0.59 \approx -113.9$ kJ mol^{-1}. [1]

Q13. (3 marks) Answer: (i) 1930 C; (ii) 0.635 g; (iii) 1930 s.
(i) $Q = It = 2.0 \times 965 = 1930$ C $= 0.020$ F. [1]
(ii) $Cu^{2+} + 2e^- \rightarrow Cu$: 0.020 F deposits 0.010 mol, i.e. 0.635 g. [1]
(iii) 1.27 g = 0.020 mol needs 0.040 F = 3860 C; t = 3860/2.0 = 1930 s. [1]

Q14. (3 marks) Answer: 30 S cm² mol⁻¹; 0.075; ≈ 1.2 × 10⁻⁵.

$\Lambda_m = \kappa \times 1000/c = 6.0 \times 10^{-5} \times 1000/0.002 = 30$ S cm² mol⁻¹. [1]

$\alpha = \Lambda_m/\Lambda^\circ_m = 30/400 = 0.075$. [1]

$K_a = c\alpha^2/(1 - \alpha) = 0.002 \times 0.005625/0.925 \approx 1.2 \times 10^{-5}$. [1]

Q15. (4 marks) Answer: (I) Cl_2; (II) NaOH forms; (III) water is reduced in preference to Na^+, or Na and Cl_2 from molten NaCl.

(I) Chlorine. [1]

(II) H^+ from water is reduced at the cathode, leaving OH^- ions, which with Na^+ make NaOH. [1]

(III) Cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$; anode: $2Cl^- \rightarrow Cl_2 + 2e^-$. [1] Water has a much higher (less negative) reduction potential than Na^+ ($E^\circ = -2.71$ V), so water is reduced in preference to Na^+. [1]

OR (III) Sodium at the cathode and chlorine at the anode. [1] Cathode: $Na^+ + e^- \rightarrow Na$; anode: $2Cl^- \rightarrow Cl_2 + 2e^-$. [1]

Q16. (5 marks) Answer: (a) $E = E^\circ - (0.059/2) \log([Ni^{2+}]/[Ag^+]^2)$; ≈ 0.96 V; (b) 110 S cm² mol⁻¹; 200 Ω.

(a) (i) A zinc rod in $ZnSO_4$ solution (anode) and a copper rod in $CuSO_4$ solution (cathode), joined by a wire and a salt bridge. Anode: $Zn \rightarrow Zn^{2+} + 2e^-$; cathode: $Cu^{2+} + 2e^- \rightarrow Cu$. [1] The salt bridge completes the circuit and keeps the two solutions electrically neutral. [1]

(ii) $Ni + 2Ag^+ \rightarrow Ni^{2+} + 2Ag$, n = 2: $E_{cell} = E^\circ_{cell} - (0.059/2) \log([Ni^{2+}]/[Ag^+]^2)$. [1]

(iii) $E^\circ_{cell} = 0.80 - (-0.25) = 1.05$ V. [½] $[Ni^{2+}]/[Ag^+]^2 = 0.1/10^{-4} = 10^3$: $E = 1.05 - 0.0295 \times 3 \approx 0.96$ V. [1½]

OR (b) (i) Conductivity κ: the conductance of a unit cube of the solution (S cm⁻¹). Molar conductivity $\Lambda_m = \kappa/c$: the conductance of all the electrolyte in one mole held between electrodes a unit distance apart. [1] On dilution κ decreases (fewer ions per unit volume) but Λ_m increases. [1] Strong electrolytes: Λ_m rises slowly and linearly with $\sqrt{c}$ ($\Lambda_m = \Lambda^\circ_m - A\sqrt{c}$), as interionic attractions fall. Weak electrolytes: Λ_m rises steeply at low concentration, because the degree of dissociation rises. [1]

(ii) $\Lambda_m = 5.5 \times 10^{-3} \times 1000/0.05 = 110$ S cm² mol⁻¹. [1] $R = G^*/\kappa = 1.1/5.5 \times 10^{-3} = 200$ Ω. [1]

Q17. (5 marks) Answer: (a) $PbSO_4$ forms at both electrodes; $H_2 + O_2 \rightarrow H_2O$; (b) ≈ 0.318 g Cu; 965 C; 1930 s.

(a) (i) A primary battery cannot be recharged once used up (dry cell, mercury cell); a secondary battery can be recharged by passing current through it in the opposite direction (lead storage battery, nickel-cadmium cell). [1] Anode: $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$. [1] Cathode: $PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$. [1]

(ii) Anode: $2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$; cathode: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$. [1] Advantages: high efficiency (about 70 %), no pollution (only water formed), works as long as the fuel is supplied. [1]

OR (b) (i) First law: the mass of substance deposited or liberated at an electrode is proportional to the charge passed ($m = ZIt$). [1] Second law: for the same charge, the masses of different substances are proportional to their equivalent masses (molar mass/number of electrons). [1]

(ii) n(Ag) = 1.08/108 = 0.010 mol, so the charge = 0.010 F = 965 C. [1] Cu: 0.010 F deposits 0.0050 mol = 0.0050 × 63.5 ≈ 0.318 g. [1] t = 965/0.50 = 1930 s. [1]

All questions are original and © Master Testing Agency. Free to download and use for personal practice; not for resale. Every numerical answer in this paper was checked independently by computer: emfs from the Nernst equation written from the balanced cell reaction, charges and masses from Faraday's laws with exact fractions.