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Redox Reactions and Electrochemistry — JEE (Main) Pattern Topic Test

FREE SAMPLE TEST · CHEMISTRY · PATTERN OF JEE (MAIN)

DURATION	MAXIMUM MARKS	QUESTIONS	NAME / ROLL NO.
60 minutes	100	25 (20 + 5)	

GENERAL INSTRUCTIONS

1. This question paper has two sections, A and B. All 25 questions are compulsory.
2. Section A has TWENTY (20) multiple-choice questions. Each has FOUR options, (1), (2), (3) and (4), of which ONLY ONE is correct.
3. Section B has FIVE (05) questions whose answer is a numerical value. Enter the value rounded off to the nearest integer.
4. Marking scheme for both sections: +4 for a correct answer, −1 for an incorrect answer, 0 if the question is not attempted.
5. Take 2.303RT/F = 0.059 V at 298 K and F = 96 500 C/mol. Use of a calculator, log tables or any other aid is not allowed.
6. The answer key and worked solutions are at the end of this booklet. Attempt the paper in 60 minutes before you look at them.

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Topics covered: Redox reactions · Oxidation number · Balancing redox reactions · Conductance of electrolytic solutions · Molar conductivity · Kohlrausch's law · Electrochemical cells · Types of electrodes · Electrode potential · Electrochemical series · Nernst equation · Cell potential and Gibbs energy · Batteries · Fuel cells

SECTION A (Q1–20) · Multiple choice · +4 / 0 / −1

Each question has FOUR options. ONLY ONE of them is correct.

1. The average oxidation number of sulphur in sodium tetrathionate, $Na_2S_4O_6$, is

(1) +2

(2) +2.5

(3) +3

(4) +5

2. The oxidation numbers of phosphorus in H_3PO_2, H_3PO_3 and H_3PO_4 are, respectively,

(1) +3, +3, +5

(2) +1, +3, +5

(3) +1, +2, +5

(4) +3, +4, +5

3. When the equation $MnO_4^- + C_2O_4^{2-} + H^+ \rightarrow Mn^{2+} + CO_2 + H_2O$ is balanced with the smallest whole-number coefficients, the coefficient of H^+ is

(1) 8

(2) 5

(3) 16

(4) 2

4. Which of the following is a disproportionation reaction?

(1) $2KClO_3 \rightarrow 2KCl + 3O_2$

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

(3) $NH_4NO_2 \rightarrow N_2 + 2H_2O$

(4) $Cl_2 + 2OH^- \rightarrow Cl^- + ClO^- + H_2O$

5. The limiting molar conductivities (in S cm² mol⁻¹) of HCl, HCOONa and NaCl are 425, 105 and 126. The limiting molar conductivity of formic acid is

(1) 404 S cm² mol⁻¹

(2) 446 S cm² mol⁻¹

(3) 656 S cm² mol⁻¹

(4) 194 S cm² mol⁻¹

6. The molar conductivity of 0.010 M formic acid is 40.4 S cm² mol⁻¹ and its limiting molar conductivity is 404 S cm² mol⁻¹. The dissociation constant of formic acid is about

(1) 1.0×10^{-4}

(2) 1.0×10^{-3}

(3) 1.1×10^{-3}

(4) 1.1×10^{-4}

7. A conductivity cell filled with a KCl solution of conductivity 1.40 S/m has a resistance of 50 Ω. Filled with a 0.050 M solution of an electrolyte, the same cell has a resistance of 175 Ω. The molar conductivity of the electrolyte is

(1) 8.0 S cm² mol⁻¹

(2) 80 S cm² mol⁻¹

(3) 280 S cm² mol⁻¹

(4) 40 S cm² mol⁻¹

8. Given $E^\circ(Cu^{2+}/Cu) = +0.34$ V and $E^\circ(Cu^{+}/Cu) = +0.52$ V, the value of $E^\circ(Cu^{2+}/Cu^{+})$ is

(1) −0.18 V

(2) +0.86 V

(3) +0.16 V

(4) +0.18 V

9. The emf at 298 K of the cell $Mg \mid Mg^{2+}(0.010\ M) \parallel Ag^{+}(0.0010\ M) \mid Ag$, for which $E^\circ(cell) = 3.17$ V, is about

(1) 3.288 V

(2) 3.052 V

(3) 3.170 V

(4) 3.111 V

10. For the cell reaction $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$, $E^\circ(cell) = 0.46$ V. The standard Gibbs energy change of the reaction is about (F = 96 500 C/mol)

(1) −88.8 kJ/mol

(2) −44.4 kJ/mol

(3) +88.8 kJ/mol

(4) −177.6 kJ/mol

11. For a cell reaction in which three electrons are transferred, the equilibrium constant at 298 K is 1.0×10^{12}. The standard emf of the cell is (take $2.303RT/F = 0.059$ V)

(1) 0.079 V

(2) 0.708 V

(3) 0.236 V

(4) 0.354 V

12. The potential at 298 K of a hydrogen electrode (H_2 gas at 1 bar) dipped in a buffer solution of pH 4.5 is

(1) +0.266 V

(2) −0.133 V

(3) −0.059 V

(4) −0.266 V

13. During the discharge of a lead storage battery, the reaction at the anode is

(1) $PbO_2 + SO_4^{2-} + 4H^{+} + 2e^{-} \rightarrow PbSO_4 + 2H_2O$

(2) $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^{-}$

(3) $PbSO_4 + 2e^{-} \rightarrow Pb + SO_4^{2-}$

(4) $Pb^{2+} + 2e^{-} \rightarrow Pb$

14. In a hydrogen–oxygen fuel cell with an alkaline (KOH) electrolyte, the reaction at the cathode is

(1) $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

(2) $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$

(3) $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

(4) $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

15. Assertion A: The molar conductivity of a weak electrolyte such as acetic acid rises steeply on dilution, especially at low concentrations.
Reason R: The degree of dissociation of a weak electrolyte increases on dilution.
In the light of the above statements, choose the correct answer from the options given below:

(1) Both A and R are true but R is NOT the correct explanation of A

(2) A is true but R is false

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

(4) A is false but R is true

16. E° values (V): Zn^{2+}/Zn −0.76, Cu^{2+}/Cu +0.34, Ag^+/Ag +0.80, Fe^{3+}/Fe^{2+} +0.77, Au^{3+}/Au +1.50. A solution of which of the following can be kept in a copper vessel without reacting with it?

(1) $AgNO_3$

(2) $Fe_2(SO_4)_3$

(3) $ZnSO_4$

(4) $AuCl_3$

17. The emf at 298 K of the concentration cell $Ag \mid Ag^+(0.0020\ M) \parallel Ag^+(0.20\ M) \mid Ag$ is

(1) 0.059 V

(2) 0 V

(3) 0.236 V

(4) 0.118 V

18. Consider the following statements.
(A) The potential of the standard hydrogen electrode is taken as zero at all temperatures.
(B) In a galvanic cell, oxidation takes place at the anode, which is the negative electrode.
(C) In an electrolytic cell, the anode is the negative electrode.
(D) A salt bridge completes the circuit and keeps the two solutions electrically neutral.
(E) The emf of a cell doubles when the cell reaction is written with all coefficients doubled.
Choose the correct answer from the options given below:

(1) (A), (B) and (D) only

(2) (A), (C) and (D) only

(3) (B), (D) and (E) only

(4) (A), (B), (D) and (E) only

19. The volume of 0.020 M $KMnO_4$ solution needed to oxidise 20 mL of 0.10 M $FeSO_4$ solution in the presence of dilute sulphuric acid is

(1) 100 mL

(2) 4 mL

(3) 20 mL

(4) 40 mL

20. In which of the following compounds does nitrogen have an oxidation number of +3?

(1) N_2O

(2) NO_2

(3) N_2O_3

(4) NH_2OH

SECTION B (Q21–25) · Numerical value · +4 / 0 / −1

The answer to each question is a NUMERICAL VALUE. Enter it rounded off to the nearest integer.

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21. The equation $Cr_2O_7^{2-} + Fe^{2+} + H^+ \rightarrow Cr^{3+} + Fe^{3+} + H_2O$ is balanced with the smallest whole-number coefficients. The sum of all the coefficients is _____.

Answer: _____

22. The electrode potential (in mV) of a copper electrode in 0.010 M Cu^{2+} solution at 298 K is _____ ($E°(Cu^{2+}/Cu)$ = +0.34 V).

Answer: _____

23. The limiting ionic conductivities of Ba^{2+} and Cl^- are 127.2 and 76.3 S cm² mol^{-1}. The limiting molar conductivity of $BaCl_2$ (in S cm² mol^{-1}) is _____ (to the nearest integer).

Answer: _____

24. For $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$, $E°(cell)$ = 0.78 V at 298 K. The value of log K for the reaction is _____ (to the nearest integer).

Answer: _____

25. The number of species among the following in which the oxidation number of the element written first (other than H, Na or K) is +6 is _____.

SO_3, H_2SO_4, $K_2Cr_2O_7$, CrO_5, $H_2S_2O_8$, $KMnO_4$, XeF_4, K_2MnO_4, SO_2Cl_2, K_2FeO_4

Answer: _____

END OF THE QUESTION PAPER

ANSWER KEY

Q1	Q2	Q3	Q4	Q5	Q6	Q7	Q8	Q9	Q10
(2)	(2)	(3)	(4)	(1)	(4)	(2)	(3)	(2)	(1)

Q11	Q12	Q13	Q14	Q15	Q16	Q17	Q18	Q19	Q20
(3)	(4)	(2)	(1)	(3)	(3)	(4)	(1)	(3)	(3)

Q21	Q22	Q23	Q24	Q25
36	281	280	26	8

WORKED SOLUTIONS

Q1. Answer (2). $2(+1) + 4x + 6(-2) = 0$ gives $x = +2.5$. (In the structure the two end S atoms are +5 and the two bridging S atoms 0; the average is +2.5.)

Q2. Answer (2). With H +1 and O −2: H_3PO_2: $3 + x - 4 = 0$, $x = +1$; H_3PO_3: $3 + x - 6 = 0$, $x = +3$; H_3PO_4: $3 + x - 8 = 0$, $x = +5$.

Q3. Answer (3). Mn goes from +7 to +2 (gains $5e^-$); each $C_2O_4^{2-}$ gives $2CO_2$ (C from +3 to +4, loses $2e^-$). Electrons balance with $2MnO_4^-$ and $5C_2O_4^{2-}$: $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$. Charge: $-2 - 10 + 16 = +4 = 2 \times (+2)$.

Q4. Answer (4). In disproportionation one element in one oxidation state is both oxidised and reduced. Chlorine (0) goes to −1 in Cl^- and +1 in ClO^-. In $KClO_3$ decomposition Cl is reduced and O oxidised (two elements); in NH_4NO_2, N at −3 and +3 meet at 0 (the reverse, comproportionation).

Q5. Answer (1). By Kohlrausch's law of independent migration, $\Lambda°(HCOOH) = \Lambda°(HCl) + \Lambda°(HCOONa) - \Lambda°(NaCl) = 425 + 105 - 126 = 404$ S cm² mol⁻¹.

Q6. Answer (4). $\alpha = \Lambda_m/\Lambda°_m = 0.10$. $K_a = c\alpha^2/(1 - \alpha) = 0.010 \times 0.01/0.90 \approx 1.1 \times 10^{-4}$. (Dropping the $1 - \alpha$ gives 1.0×10^{-4}.)

Q7. Answer (2). Cell constant $G^* = \kappa R = 1.40 \times 50 = 70\ m^{-1}$. κ of the solution = 70/175 = 0.40 S/m. $\Lambda_m = \kappa/c = 0.40\ S/m \div 50\ mol/m^3 = 8.0 \times 10^{-3}$ S m² mol⁻¹ = 80 S cm² mol⁻¹.

Q8. Answer (3). Potentials are not additive, but $\Delta G° = -nFE°$ is. $Cu^{2+} + 2e^- \rightarrow Cu$: $-2F(0.34)$. Subtract $Cu^+ + e^- \rightarrow Cu$: $-1F(0.52)$. This leaves $Cu^{2+} + e^- \rightarrow Cu^+$ with $-FE° = -0.68F + 0.52F$, so $E° = +0.16$ V. (Subtracting the potentials directly gives the wrong −0.18 V.)

Q9. Answer (2). $Mg + 2Ag^+ \rightarrow Mg^{2+} + 2Ag$, n = 2. $Q = [Mg^{2+}]/[Ag^+]^2 = 0.010/10^{-6} = 10^4$. $E = 3.17 - (0.059/2) \times 4 = 3.17 - 0.118 = 3.052$ V. (Forgetting to square $[Ag^+]$ gives 3.111 V.)

Q10. Answer (1). Two electrons pass per formula of the reaction: $\Delta G° = -nFE° = -2 \times 96\,500 \times 0.46$ J/mol ≈ −88.8 kJ/mol.

Q11. Answer (3). At equilibrium E = 0, so $E° = (0.059/n) \log K = (0.059/3) \times 12 = 0.236$ V. (Taking n = 2 gives 0.354 V; forgetting n gives 0.708 V.)

Q12. Answer (4). $2H^+ + 2e^- \rightarrow H_2$: $E = 0 - (0.059/2) \log(p(H_2)/[H^+]^2) = 0.059 \log[H^+] = -0.059 \times 4.5 \approx -0.266$ V.

Q13. Answer (2). On discharge the lead plate is the anode, where oxidation takes place: $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$. PbO_2 is reduced to $PbSO_4$ at the cathode. The third reaction happens at the same plate only while charging.

Q14. Answer (1). Oxygen is reduced at the cathode; in alkaline solution it gives hydroxide: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$. Hydrogen is oxidised at the anode (the second equation). The overall reaction is $2H_2 + O_2 \rightarrow 2H_2O$.

Q15. Answer (3). $\Lambda_m = \alpha\Lambda^\circ_m$ for a weak electrolyte, and α grows on dilution (Ostwald: $\alpha \approx \sqrt{(K_a/c)}$), so Λ_m rises steeply as $c \to 0$. Both are true and R explains A.

Q16. Answer (3). Copper is oxidised by any ion whose reduction potential is above +0.34 V: Ag^+ (0.80), Fe^{3+} (0.77) and Au^{3+} (1.50). Zn^{2+} (−0.76) cannot oxidise copper, so $ZnSO_4$ solution is safe.

Q17. Answer (4). $E^\circ = 0$ for a concentration cell. With n = 1, $E = 0.059 \log(0.20/0.0020) = 0.059 \times 2 = 0.118$ V; the electrode in the more concentrated solution is the cathode.

Q18. Answer (1). (A), (B) and (D) are true. (C) is false: in electrolysis the anode is connected to the positive terminal. (E) is false: emf is intensive; doubling the equation doubles n and ΔG° but $E^\circ = -\Delta G^\circ/nF$ stays the same.

Q19. Answer (3). $MnO_4^- + 5Fe^{2+} + 8H^+ \to Mn^{2+} + 5Fe^{3+} + 4H_2O$. Fe^{2+} = 2.0 mmol needs 0.40 mmol MnO_4^-: V = 0.40/0.020 = 20 mL.

Q20. Answer (3). N_2O: +1; NO_2: +4; N_2O_3: 2x = 6, x = +3; NH_2OH: x + 3 − 2 = 0, x = −1.

Q21. Answer 36. $Cr_2O_7^{2-}$ gains $6e^-$ (2Cr from +6 to +3); each Fe^{2+} loses $1e^-$. $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \to 2Cr^{3+} + 6Fe^{3+} + 7H_2O$. Sum = 1 + 6 + 14 + 2 + 6 + 7 = 36.

Q22. Answer 281. $Cu^{2+} + 2e^- \to Cu$: $E = 0.34 - (0.059/2) \log(1/0.010) = 0.34 - 0.059 = 0.281$ V = 281 mV.

Q23. Answer 280. $\Lambda^\circ(BaCl_2) = \lambda^\circ(Ba^{2+}) + 2\lambda^\circ(Cl^-) = 127.2 + 152.6 = 279.8$, i.e. 280.

Q24. Answer 26. $\log K = nE^\circ/0.059 = 2 \times 0.78/0.059 \approx 26.4$, i.e. 26 ($K \approx 3 \times 10^{26}$, which is why an iron nail gets coated with copper).

Q25. Answer 8. +6: S in SO_3, H_2SO_4 and SO_2Cl_2; Cr in $K_2Cr_2O_7$; Cr in CrO_5 (one oxide O and four peroxide O: x = 2 + 4 = 6); S in $H_2S_2O_8$ (two peroxide O: 2x = 14 + 2 − 2 − 2 = 12); Mn in K_2MnO_4; Fe in K_2FeO_4. $KMnO_4$ is +7 and XeF_4 +4. Eight in all.

All questions are original and © Master Testing Agency. Free to download and use for personal practice; not for resale. Every answer in this paper was checked independently by computer: oxidation numbers were solved from the formulae, redox equations balanced exactly so that every atom and the charge are conserved, cell potentials found from the Nernst equation and $\Delta G = -nFE$, and conductivities from their definitions.