Electrode potential, E_{Cell}, Nernt equation and ECS

51. Amongst the following electrodes the one with zero electrode potential is
(a) Calomel electrode
(b) Standard hydrogen electrode
(c) Glass electrode
(d) Gas electrode

52. Which of the following is correct expression for electrode potential of a cell
(a) $E = E^o - \frac{RT}{nF} ln \frac{[product]}{[reactant]}$
(b) $E = E^o + \frac{RT}{F} ln \frac{[product]}{[reactant]}$
(c) $E = E^o - \frac{RT}{nF} ln \frac{[reactant]}{[product]}$
(d) $E = -\frac{RT}{F} ln \frac{[product]}{[reactant]}$

53. Calculate standard free energy change for the reaction $\frac{1}{2}Cu(s) + \frac{1}{2}Cl_2(g) \rightleftharpoons \frac{1}{2}Cu^{2+} + Cl^-$ taking place at 25^oC in a cell whose standard e.m.f. is 1.02 *volts*
(a) – 98430 *J* (b) 98430 *J*
(c) 96500 *J* (d) – 49215 *J*

54. In which cell the free energy of a chemical reaction is directly converted into electricity ?
(a) Leclanche cell
(b) Concentration cell
(c) Fuel cell
(d) Lead storage battery

55. Nernst equation is related with
(a) The electrode potential and concentration of ions in the solution
(b) Equilibrium constant and concentration of ions
(c) Free energy change and E.M.F. of the cell
(d) None of these

56. The standard reduction potentials of 4 elements are given below. Which of the following will be the most suitable reducing agent
I = – 3.04 *V*, II = – 1.90 V, III = 0 *V*, IV = 1.90 *V*
(a) I (b) II
(c) III (d) IV

57. Electrode potential data are given below :
$Fe^{3+}(aq) + e^- \rightarrow Fe^{-1}(aq); E^o = +0.77V$
$Al^{3+}(aq) + 3e^- \rightarrow Al(s); E^o = -1.66V$
$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq); E^o = +1.08V$

Based on the data given above, reducing power of Fe^{2+}, Al and Br^- will increase in the order

(a) $Br^- < Fe^{2+} < Al$

(b) $Fe^{2+} < Al < Br^-$

(c) $Al < Br^- < Fe^{2+}$

(d) $Al < Fe^{2+} < Br^-$

58. The standard electrode potential (E^o) for OCl^-/Cl^- and $Cl^-/\frac{1}{2}Cl_2$ respectively are $0.94V$ and $-1.36V$. The E^o value for $OCl^-/\frac{1}{2}Cl_2$ will be

(a) $-0.42V$ (b) $-2.20V$

(c) $0.52V$ (d) $1.04V$

59. If the reduction potential is more, then

(a) It is easily oxidised

(b) It is easily reduced

(c) It acts as oxidising agent

(d) It has redox nature

60. One of the following is false for Hg

(a) It can evolve hydrogen from H_2S

(b) It is a metal

(c) It has high specific heat

(d) It is less reactive than hydrogen

61. E^o for the cell $Zn|Zn^{2+}(aq)||Cu^{2+}(aq)|Cu$ is $1.10V$ at 25^oC, the equilibrium constant for the reaction $Zn + Cu^{2+}(aq) \rightleftharpoons Cu + Zn^{2+}(aq)$ is of the order of

(a) 10^{-28} (b) 10^{-37}

(c) 10^{+18} (d) 10^{+17}

62. Standard reduction potentials at 25^oC of $Li^+|Li, Ba^{2+}|Ba, Na^+|Na$ and $Mg^{2+}|Mg$ are $-3.05, -2.90, -2.71$ and -2.37 *volt* respectively. Which one of the following is the strongest oxidising agent

(a) Na^+ (b) Li^+

(c) Ba^{2+} (d) Mg^{2+}

63. Which of the following displaces Br_2 from an aqueous solution containing bromide ions

(a) Cl_2 (b) Cl^-

(c) I_2 (d) I_3^-

64. For the cell reaction

$Cu^{2+}(C_1aq) + Zn(s) = Zn^{2+}(C_2aq) + Cu(s)$

of an electrochemical cell, the change in free energy at a given temperature is a function of

(a) ln (C_1) (b) ln (C_2)

(c) ln $(C_1 + C_2)$ (d) ln (C_2/C_1)

65. The e.m.f. of the cell in which the following reaction $Zn(s) + Ni^{2+}(a = 1.0) \rightleftharpoons Zn^{2+}(a = 10) + Ni(s)$ occurs, is found to be $0.5105V$ at $298 \rightleftarrows K$. The standard e.m.f. of the cell is

(a) 0.5400 (b) 0.4810 *V*
(c) 0.5696 *V* (d) – 0.5105 *V*

66. For the redox reaction $Zn(s) + Cu^{2+}(0.1M) \rightarrow Zn^{2+}(1M) + Cu(s)$ taking place in a cell, E^o_{cell} is 1.10 *volt*. E_{cell} for the cell will be $\left(2.303\frac{RT}{F} = 0.0591\right)$

(a) 2.14 *volt* (b) 1.80 *volt*
(c) 1.07 *volt* (d) 0.82 *volt*

67. The *emf* of a Daniel cell at 298*K* is E_1 $Zn|\underset{(0.01M)}{ZnSO_4}||\underset{(1.0M)}{CuSO_4}|Cu$ when the concentration of $ZnSO_4$ is 1.0 *M* and that of $CuSO_4$ is 0.01 *M*, the *emf* changed to E_2. What is the relationship between E_1 and E_2

(a) $E_2 = 0 \neq E_1$ (b) $E_1 > E_2$
(c) $E_1 < E_2$ (d) $E_1 = E_2$

68. The oxidation potentials of following half-cell reactions are given $Zn \rightarrow Zn^{2+} + 2e^-; E^o = 0.76V$, $Fe \rightarrow Fe^{2+} + 2e^-; E^o = 0.44V$ what will be the *emf* of cell, whose cell-reaction is $Fe^{2+}(aq) + Zn \rightarrow Zn^{2+}(aq) + Fe$

(a) – 1.20 *V* (b) + 0.32 *V*
(c) – 0.32 *V* (d) + 1.20 *V*

69. The E^o for half cells Fe/Fe^{2+} and Cu/Cu^{2+} are – 0.44 *V* and + 0.32 *V* respectively. Then

(a) Cu^{2+}oxidises Fe
(b) Cu^{2+}oxidises Fe^{2+}
(c) Cu oxidises Fe^{2+}
(d) Cu reduces Fe^{2+}

70. What is E^o for electrode represented by $Pt, O_2(1atm)/2H^+(Im)$

(a) Unpredictable (b) Zero
(c) 0.018 *V* (d) 0.118 *V*

71. The cell potential of a cell in operation is

(a) Zero
(b) Positive
(c) Negative
(d) None of the above

72. Which of the following is displaced by Fe

(a) Ag (b) Hg
(c) Zn (d) Na

73. The standard electrode potential of the half cells are given below $Zn^{2+} + 2e^- \rightarrow Zn; E = -7.62V, Fe^{2+} + 2e^- \rightarrow Fe; E = -7.81V$

The *emf* of the cell $Fe^{2+} + Zn \rightarrow Zn^{2+} + Fe$ is

(a) 1.54 *V* (b) – 1.54 *V*
(c) – 0.19 *V* (d) + 0.19 *V*

74. $Zn^{2+} + 2e^- \rightarrow Zn(s); E^o = -0.76$,
$Fe^{3+} + e^- \rightarrow Fe^{2+}; E^o =$
$-0.77, Cr^{3+} + 3e^- \rightarrow Cr; E^o = -0.79,$
$H^+ + 2e^- \rightarrow 1/2H_2; E^o = 0.00$
Strongest reducing agent is
(a) Fe^{2+} (b) Zn
(c) Cr (d) H_2

75. Standard reduction electrode potentials of three metals *A*, *B* and *C* are respectively + 0.5*V*, – 3.0*V* and – 1.2 *V*. The reducing powers of these metals are
(a) *B* > *C* > *A* (b) *A* > *B* > *C*
(c) *C* > *B* > *A* (d) *A* > *C* > *B*