Important Questions

Multiple Choice questions-

Question 1. If the conductivity and conductance of a solution is same then its cell constant is equal to:

(a) 1

(b) 0

(c) 10

(d) 1000

Question 2.The units of conductivity are:

(a) ohm^{-1}

(b) $ohm^{-1}\ cm^{-1}$

(c) $ohm^{-2}\ cm^2\ equiv^{-1}$

(d) $ohm^{-1}\ cm^2$

Question 3.The resistance of 0.1 N solution of acetic acid is 250 ohm, when measured in a cell of cell constant 1.15 cm^{-1}. The equivalent conductance (in $ohm^{-1}\ cm^2\ equivalent^{-1}$) of 0.1 N acetic acid is

(a) 18.4

(b) 0.023

(c) 46

(d) 9.2

Question 4.In infinite dilution of aqueous solution of BaCl2, molar conductivity of Ba^{2+} and Cl^- ions are = 127.32 S cm^2/mol and 76.34 S cm^2/mol respectively. What is A°m for BaCl2 at same dilution?

(a) 280 S $cm^2\ mol^{-1}$

(b) 330.98 S cm^2 mol^{-1}

(c) 90.98 S cm^2 mol^{-1}

(d) 203.6 S cm^2 mol^{-1}

Question 5.The specific conductance of 0.1 M NaCl solution is 1.06×10^{-2} ohm^{-1} cm^{-1}. Its molar conductance in ohm^{-1} cm^2 mol^{-1} is

(a) 1.06×10^2

(b)1.06×10^3

(c) 1.06×10^4

(d) 53

Question 6.The limiting molar conductivities A° for NaCl, KBr and KCl are 126, 152 and 150 S cm^2 mol^{-1} respectively. The A° for NaBr is

(a) 278 S cm^2 mol^{-1}

(b) 976 S cm^2 mol^{-1}

(c) 128 S cm^2 mol^{-1}

(d) 302 S cm^2 mol^{-1}

Question 7. λ(CICH2COONa) = 224 ohm^{-1} cm^2 gm eq^{-1}, λ(NaCl) = 38.2 ohm^{-1} cm^2 gm eq^{-1}. λ(HCl) = 203 ohm^{-1} cm^2 gm eq^{-1}, what is the value of λ(CICH2COOH)?

(a) 288.5 ohm^{-1} cm^2 gm eq^{-1}

(b) 289.5 ohm^{-1} cm^2 gm eq^{-1}

(c) 388.8 ohm^{-1} cm^2 gm eq^{-1}

(d) 59.5 ohm^{-1} cm^2 gm eq^{-1}

Question 8. The limiting molar conductivities of HCl, CH_3COONa and NaCl are respectively 425, 90 and 125 mho cm² mol^{-1} at 25 °C. The molar conductivity of 0.1 M CH3COOH solution is 7.8 mho cm² mol^{-1} at the same temperature. The degree of dissociation of 0.1 M acetic acid solution at the same temperature is

(a) 0.10

(b) 0.02

(c) 0.15

(d) 0.03

Question 9.The values of limiting ionic conductance of H and HCOO– ions are respectively 347 and 53 S cm² mol^{-1} at 298 K. If the molar conductance of 0.025 M methanoic acid at 298 K is 40 S cm² mol^{-1}, the dissociation constant of methanoic acid at 298 K is

(a) 1×10^{-5}

(b) 2×10^{-5}

(c) 1.5 × 10-^{4}

(d) 2.5×10^{-4}

Question 10.The ionisation constant of a weak electrolyte is 2.5×10^{-5} and molar conductance of its 0.01 M solution is 19.6 S cm² mol^{-1}. The molar conductance at infinite dilution (S cm² mol^{-1}) is

(a) 402

(b) 392

(c) 306

(d) 39.2

Very Short Question:

Question 1. Can you store AgCl solution in Zinc pot?

Question 2. Define the term – standard electrode potential?

Question 3. What is electromotive force of a cell?

Question 4. Can an electrochemical cell act as electrolytic cell? How?

Question 5. Single electrode potential cannot be determined. Why?

Question 6. What is SHE? What is its electrode potential?

Question 7. What does the positive value of standard electrode potential indicate?

Question 8. What is an electrochemical series? How does it predict the feasibility of a certain redox reaction?

Question 9. Give some uses of electrochemical cells?

Question 10. State the factors that affect the value of electrode potential?

Short Questions:

Question 1. What is the cell potential for the cell at $25^0C \quad Cr/Cr^{3+}10.1\ m]//Fe^{2+}(0.01m)/Fe$

$E^0_{cr+/cr} = -0.74V$. $E^0Fe^{2+}/Fe = -0.44V$

Question 2. Calculate ΔG^0 for the reaction

$25^0C \; Zn(s)1Zn^2+[0.0004m]11cd^2+(0.2m)1cd(s) \; E^0{}_{Zn}{}^{2+}{}_{/Zn} = -0.763V$

$E^0{}_{cd}{}^{2+}/cd = -0.403v, \; F = 96500\ CMol^{-1}$ R = 8.314 J/K

Question 3. Calculate Equilibrium constant K for the reaction

$298K\ Zn(s)+Cu^{2+}(aq) \rightleftharpoons Zn^{2+}/aq)+Cu\ E^0{}_{Zn}{}^{2+}1Zn = -0.076v; E^0Cu^2+/Cu$ +0.34v.

Question 4. For what concentration of $Ag+(aq)$ will the emf of the given cell be zero at 25^0C if the concentration of $CU^{2+}(aq)$ is 0.1 M? $Cu(s)/Cu^{2+}(0.1M)//Ag^+(aq)/Ag(s)$
$E^0Ag^+/Ag = +0.80V; \; E^0{}_{Cu}{}^{2+/Cu} = 0.34\ V$

Question 5. Calculate the standard free energy change for the cell- reaction.

$Fe^{2+}(aq)+Ag^+(s)\grave{a}\rightarrow Fe^3+(aq)+Ag(s)$ How is it related to the equilibrium

constant of the reaction? $E^0{}_{Fe}{}^{3+}/Fe^{2+} = +0.77V$, $E^0{}_{Ag}{}^{+1Ag} = +0.08V \; F = 96500\ C/mol.$

Question 6. How much charge is required for the following reductions:

(i) 1 mol of to Al.
(ii) 1 mol of to Cu.
(iii) 1 mol of to .

Question 7. How much electricity in terms of Faraday is required to produce

(i) 20.0 g of Ca from molten .
(ii) 40.0 g of Al from molten

Question 8. How much electricity is required in coulomb for the oxidation of

(i) 1 mol of to .
(ii) 1 mol of FeO to

Question 9. A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Question 10. Depict the galvanic cell in which the reaction takes place. Further show:

(i) Which of the electrode is negatively charged?
(ii) The carriers of the current in the cell.
(iii) Individual reaction at each electrode.

Long Questions:

Question 1. Explain construction and working of standard Hydrogen electrode? (b) Write any two differences between amorphous solids and crystalline solids.

Question 2.

The molar conductivity of 0.025 mol L-1 methanoic acid is 46.1 S $cm^2\,mol^{-1}$. Calculate its degree of dissociation and dissociation constant. Given $\lambda°H^+$= 349.6 S $cm^2\,mol^{-1}$ and $\lambda°(HCOO-) = 54.6\,S\,cm^2\,mol$

Question 3.Explain how rusting of iron is envisaged as setting up of an electrochemical cell.

Question 4 Calculate the standard cell potentials of galvanic cells in which the following reactions take place:

Question 4. Write the Nernst equation and emf of the following cells at 298 K:

Question 5. Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

Assertion and Reason Questions:

1. In these questions, a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

a) Assertion and reason both are correct statements and reason is correct explanation for assertion.
b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.
c) Assertion is correct statement but reason is wrong statement.
d) Assertion is wrong statement but reason is correct statement.

Assertion: At the end of electrolysis using platinum electrodes, an aqueous solution of copper sulphate tums colourless.

Reason: Copper in $CuSO_4$ is converted to $Cu(OH)_2$ during the electrolysis.

2. In these questions, a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

a) Assertion and reason both are correct statements and reason is correct explanation for assertion.
b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.
c) Assertion is correct statement but reason is wrong statement.
d) Assertion is wrong statement but reason is correct statement.

Assertion: Zinc displaces copper from copper sulphate solution.

Reason: Eº of zinc is -0.vV and that of copper is +0.34V.

Case Study Questions:

1. The concentration of potassium ions inside a biological cell is at least twenty times higher than the outside. The resulting potential difference across the cell is important in several processes such as transmission of nerve impulses and maintaining the ion balance. A simple model for such a concentration cell involving a metal M is,

$M_{(s)}$ | M^+(aq.; 0.05 molar) || M^+(aq; 1 molar) | $M_{(s)}$

The following questions are multiple choice questions. Choose the most appropriate answer:

(i) For the above cell,

a) $E_{cell} < 0; \Delta G > 0$

b) $E_{cell} > 0; \Delta G < 0$
c) $E_{cell} < 0; \Delta G^{\circ} > 0$
d) $E_{cell} > 0; \Delta G^{\circ} < 0$

(ii) If the 0.05 molar solution of M^+ is replaced by a 0.0025 molar M^+ solution, then the magnitude of the cell potential would be:

a) 130mV
b) 185mV
c) 154mV
d) 600mV

(iii)The value of equilibrium constant for a feasible cell reaction is:

a) < 1
b) = 1
c) > 1
d) Zero

(iv)What is the emf of the cell when the cell reaction attains equilibrium?

a) 1
b) 0
c) > 1
d) < 1

(v) The potential of an electrode change with change in:

a) Concentration ofions in solution.
b) Position of electrodes.
c) Voltage of the cell.
d) All of these.

2. All chemical reactions involve interaction of atoms and molecules. A large number of atoms/molecules are present in a few gram of any chemical compound varying with their atomic/ molecular masses. To handle such large number conveniently, the mole concept was introduced. All electrochemical cell reactions are also based on mole concept. For example, a 4.0 molar aqueous solution of NaCl is prepared and 500mL of this solution is electrolysed. This leads to the evolution of chlorine gas at one of the electrode. The amount of products formed can be calculated by using mole concept.

The following questions are multiple choice questions. Choose the most appropriate answer:

(i) The total number of moles of chlorine gas evolved is:

a) 0.5
b) 1.0
c) 1.5
d) 1.9

(ii) If cathode is a Hg electrode, then the maximum weight of amalgam formed from this solution is:

a) 300g
b) 446g
c) 396g
d) 296g

(iii)The total charge (coulomb) required for complete electrolysis is:

a) 186000
b) 24125
c) 48296
d) 193000

(iv)In the electrolysis, the number of moles of electrons involved are:

a) 2
b) 1
c) 3
d) 4

(v) In electrolysis of aqueous NaCl solution when Pt electrode is taken, then which gas is liberated at cathode?

a) H_2gas
b) Cl_2gas
c) O_2gas
d) None of these.

Answers key

MCQ answers:

1. Answer: (a) 1

2. Answer: (b) ohm^{-1} cm^{-1}

3. Answer: (c) 46

4. Answer: (a) 280 S cm^2 mol^{-1}

5. Answer: (a) 1.06×10^2

6. Answer: (c) 128 S cm^2 mol^{-1}

7. Answer: (c) 388.8 ohm^{-1} cm^2 gm eq^{-1}

8. Answer: (b) 0.02

9. Answer: (d) 2.5×10^{-4}

10. Answer: (b) 392

Very Short Answers:

1. No. We can't store AgCl solution in Zinc pot because standard electrode potential of Zinc is less than silver..
2. When the concentration of all the species involved in a half-cell is unity, then the electrode potential is called standard electrode potential.

3. Answer: Electromotive force of a cell is also called the cell potential. It is the difference between the electrode potentials of the cathode and anode.

$$E_{cell} = E_{cathode} - E_{anode}$$

4. Answer: Yes, An electrochemical cell can be converted into electrolytic cell by applying an external opposite potential greater than its own electrical potential.

5. Answer: A single half cell does not exist independently as reduction and oxidation occur simultaneously therefore single electrode potential cannot be measured.

6. Answer: SHE stands for standard Hydrogen electrode. By convention, its electrode potential is taken as 0 (zero).

7. Answer: The positive value of standard electrode potential indicates that the element gets reduced more easily than ions and its reduced form is more stable than Hydrogen gas.

8. The arrangement of metals and ions in increasing order of their electrode potential values is known as electrochemical series.The reduction half reaction for which the reduction potential is

lower than the other will act as anode and one with greater value will act as cathode. Reverse reaction will not occur.

9. Electrochemical cells are used for determining the
    - pH of solutions
    - solubility product and equilibrium constant
    - in potentiometric titrations
10. Factors affecting electrode potential values are –
    - Concentration of electrolyte
    - Temperature.

Short Answers:

1. **Answer**

The cell reaction is

$$2Cr + 3Fe^{2+}\ 6e^- \rightarrow 2Cr^{3+} + 3Fe$$

Nernst Equation –

$$E_{cell} = \left(E^0_{Fe^2+/Fe} - E^0_{cr^{3+}/cr}\right) - \frac{0.059}{6}\log\frac{\left[Cr^{3+}\right]^2}{\left[Fe^{2+}\right]^3}$$

= (-0.44v – (-0.74v) – $\frac{0.059}{6}\log\frac{(0.10)^2}{(0.01)^3}$

= $0.3V - \frac{0.059}{6}\log 10^4$

= 0.3V – 0.0394V

= +0.2606 V

2. **Answer:**

The half-cell reactions are

$$Zn(s) \rightarrow Zn^2 + (aq) + 2e^-$$

Anode:

Cathode: $Cd^{2}+(aq)+2e^{-}\rightarrow Cd(s)$

Nernst Equation

$$E_{cell}=\left(E^{0}{}_{Cathode}-E^{0}{}_{anode}\right)-\frac{0.059}{n}\log\frac{\left[Zn^{2+}\right]}{\left[Cd^{2+}\right]}$$

= (- 0.403 – (-0.763) –

= 0.36V – 0.0798V = 0.4398V

$\Delta G^{0} = -\mathrm{n\,F\,E^{0}{}_{cell}}$

$$=\frac{-2mol\times96500\ C}{mol\times0.4398V}$$

= - 8488 J mol^{-1}

3. Answer:

From the reaction, n =2

$E^{0}{}_{cell}=E^{0}cu^{2}+/cu-E^{0}Zn^{2}+/Zn$

= + 0.34v – (-0.76v) = 1.10V

$$E^{0}{}_{cell}=\frac{2.303RT}{nF}\log \mathrm{k_c}$$

At 298k, $E^{0}{}_{cell}\times\frac{n}{0.059}\log \mathrm{k_c}$

Log k_c = $E^{0}{}_{cell}\times\frac{n}{0.059}$

$1.10\times\frac{2}{0.059}=37.29$

$K_c = Antilog\ 37.29$

$=1.95\times10^{37}$

4. Answer:

$\left[Ag^{+}\right]=5.3\times10^{-9}M$

5. Answer :

$E^0_{cell} = 0.03V$

6. Answer

(i)

Therefore, Required charge = 3 F

= 289461 C

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

Therefore, Required charge = 2 F

$= 2 \times 96487$ C

= 192974 C

(iii) $MnO_4^- \rightarrow Mn^{2+}$

i.e., $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$

Therefore, Required charge = 5 F

$= 5 \times 96487$ C

= 482435 C

7. Answer:

.(i) According to the question,

Electricity required to produce 40 g of calcium = 2 F

Therefore, electricity required to produce 20 g of calcium =

= 1 F

(ii) According to the question,

$Al^{3+} + 3e^{-1} \rightarrow \underset{27g}{Al}$

Electricity required to produce 27 g of Al = 3 F

Therefore, electricity required to produce 40 g of Al = $\frac{3 \times 40}{27}$F

= 4.44 F

8. Answer:

(i) According to the question,

Now, we can write:

$O^2 \rightarrow \frac{1}{2}O_2 + 2e^-$

Electricity required for the oxidation of 1 mol of H_2O to O_2= 2 F

$= 2 \times 96487$ C

= 192974 C

(ii) According to the question

$Fe^{2+} \rightarrow Fe^{3+} + e^{-1}$

Electricity required for the oxidation of 1 mol of FeO to Fe_2O_3 = 1 F
= 96487 C

9. Answer :

Given,
Current = 5A
Time = 20×60 = 1200 s
Therefore, $\text{Charge} = \text{current} \times \text{time}$

$= 5 \times 1200$
= 6000 C
According to the reaction,
$Ni^{2+}_{(aq)} + 2e^- \rightarrow \underset{58.7g}{Ni_{(s)}}$
Nickel deposited by 2×96487 C = 58.71 g
Therefore, nickel deposited by 6000 C $= \frac{58.71 \times 6000}{2 \times 96487} g$
= 1.825 g
Hence, 1.825 g of nickel will be deposited at the cathode.

10.Answer :

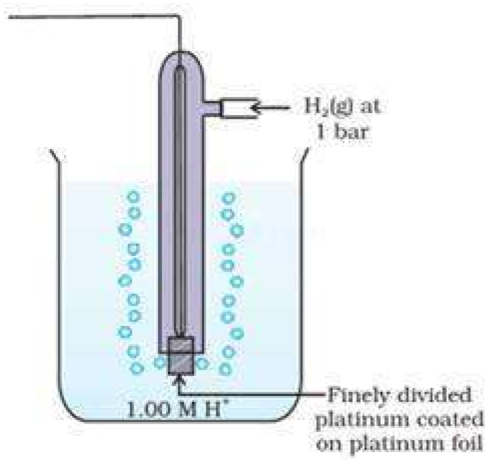
The galvanic cell in which the given reaction takes place is depicted as:
$Zn_{(s)} | Zn^{2+}_{(aq)} || Ag^{+}_{(aq)} | Ag_{(s)}$
(i) Zn electrode (anode) is negatively charged.
(ii) Ions are carriers of current in the cell and in the external circuit, current will flow from silver to zinc.
(iii) The reaction taking place at the anode is given by,
$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$
The reaction taking place at the cathode is given by,
$Ag^{+}_{(aq)} + e^- \rightarrow Ag_{(s)}$

Long Answers:

1. **Answer:**

Construction: SHE consists of a platinum electrode coated with platinum black. The electrode is dipped in an acidic solution and pure Hydrogen gas is bubbled through it. The concentration of both the reduced and oxidized. Forms of Hydrogen is maintained at unity i.e) pressure of gas is 1 bar and concentration of Hydrogen ions in the solution is 1 molar.

Working – The reaction taking place in SHE is At 298 K , the emf of the cell constructed by taking SHE as anode and other half-cell as cathode, gives the reduction potential of the other half cell whereas for a cell constructed by taking SHE as anode gives the oxidation potential of other half cell as conventionally the electrode potential of SHE is zero.



2. Answer:

$\Lambda_m = 46.1 \, S\,cm^2\,mol^{-1}$

$\lambda^\circ(H^+) = 349.6 \, S\,cm^2\,mol^{-1}$

$\lambda^\circ(HCOO^-) = 54.6 \, S\,cm^2\,mol^{-1}$

$\Lambda_m^\circ(HCOOH) = \lambda^\circ(H^+) + \lambda^\circ(HCOO^-)$

= 349.6 + 54.6 = $404.2 \, S\,cm^2\,mol^{-1}$

Now, degree of dissociation:

$$\alpha = \frac{\Lambda_m(HCOOH)}{\Lambda_m^0(HCOOH)}$$

$= \frac{46.1}{404.2}$ = 0.114 (approximately)

Thus, dissociation constant:

$$K = \frac{c\alpha^2}{(1-\alpha)}$$

$$= \frac{(0.025\,\text{mol L}^{-1})(0.114)^2}{(1-0.114)}$$

$$= 3.67 \times 10^{-4}\,\text{mol L}^{-1}$$

3. **Answer:**

In the process of corrosion, due to the presence of air and moisture, oxidation takes place at a particular spot of an object made of iron. That spot behaves as the anode. The reaction at the anode is given by,

Electrons released at the anodic spot move through the metallic object and go to another spot of the object.

There, in the presence of ions, the electrons reduce oxygen. This spot behaves as the cathode. These H+ ions come either from , which are formed due to the dissolution of carbon dioxide from air into water or from the dissolution of other acidic oxides from the atmosphere in water.

The reaction corresponding at the cathode is given

$$O_{2(g)} + 4H^+_{(aq)} + 4e^- \rightarrow 2H_2O_{(l)}$$

The overall reaction is: $2Fe_{(s)} + O_{2(g)} + 4H^+_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + 2H_2O_{(l)}$
Also, ferrous ions are further oxidized by atmospheric oxygen to ferric ions. These ferric ions combine with moisture, present in the surroundings, to form hydrated ferric oxide (Fe_2O_3, xH_2O) i.e., rust.
Hence, the rusting of iron is envisaged as the setting up of an electrochemical cell.

4. **Answer (i)** $E^\circ_{Cr^{3+}/Cr} = 0.74\text{V}$
$E^\circ_{Cd^{2+}/Cd} = 0.40\text{V}$

The galvanic cell of the given reaction is depicted as: $Cr_{(s)} \mid Cr^{3+}_{(aq)} \parallel Cd^{2+}_{(aq)} \mid Cd_{(s)}$
Now, the standard cell potential is $E^\circ_{cell} = E^\circ_R - E^\circ_L$
= 0.40 – (–0.74)
= +0.34 V
$\Delta_r G^\circ = -nFE^\circ_{cell}$
In the given equation,
n = 6
F = 96487 C mol – 1
E°_{cell} = +0.34 V
Then, $\Delta_r G^\circ = -\ 6 \times 96487\ \text{C mol}^{-1} \times 0.34\ \text{V}$
= –196833.48 CV mol^{-1}

= –196833.48 J mol^{-1}
= –196.83 kJ v
Again, $\Delta_r G^\circ = -RT \ln K$
$\Delta_r G^\circ = -2.303 RT \ln K$
$\log K = -\frac{\Delta_r G}{2.303 RT}$
$= \frac{196.83 \times 10^3}{2.303 \times 8.314 \times 298}$
= 34.496
Therefore, K = antilog (34.496)
$= 3.13 \times 1034$

(ii) $E^\circ_{Fe^{3+}/Fe^{2+}} = 0.77 V$
$E^\circ_{Ag^+/Ag} = 0.80 V$

The galvanic cell of the given reaction is depicted as: $Fe^{2+}_{(aq)} | Fe^{3+}_{(aq)} || Ag^+_{(aq)} | Ag_{(s)}$
Now, the standard cell potential is $E^\circ_{cell} = E^\circ_R - E^\circ_L$
= 0.80 – 0.77
= 0.03 V
Here, n = 1.
Then, $\Delta_r G^\circ = -nFE^\circ_{cell}$
$= -1 \times 96487 \ C \ mol^{-1} \times 0.03 \ V$
= –2894.61 J mol^{-1}
= –2.89 kJ mol^{-1}
Again, $\Delta_r G^\circ = 2.303 RT \ln K$
$\log K = -\frac{\Delta_r G}{2.303 RT}$
$= \frac{-2894.61}{2.303 \times 8.314 \times 298}$
= 0.5073
Therefore, K = antilog (0.5073)
= 3.2 (approximately)

5. Answer (i) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log \frac{[Mg^{2+}]}{[Cu^{2+}]}$$

$$= \{0.34-(-236)\} - \frac{0.0591}{2}\log\frac{.001}{.0001}$$

$$= 2.7 - \frac{0.0591}{2}\log 10$$

= 2.7 – 0.02955
= 2.67 V (approximately)

(ii) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{n}\log\frac{[Fe^{2+}]}{[H^+]^2}$$

$$= \{0-(-0.44)\} - \frac{0.0591}{2}\log\frac{0.001}{1^2}$$

= 0.44–0.02955(–3)
= 0.52865 V
= 0.53 V (approximately)

(iii) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{n}\log\frac{[Sn^{2+}]}{[H^+]^2}$$

$$= \{0-(-0.14)\} - \frac{0.0591}{2}\log\frac{0.050}{(0.020)^2}$$

$$= 0.14 - 0.0295 \times \log 125$$

= 0.14–0.062
= 0.078 V
= 0.08 V (approximately)

(iv) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{n}\log\frac{1}{[Br^-]^2[H^+]^2}$$

$$= (0-1.09) - \frac{0.0591}{2}\log\frac{1}{(0.010)^2(0.030)^2}$$

$$= -1.09 - 0.02955 \times \log\frac{1}{0.00000009}$$

$$= -1.09 - 0.02955 \times \log\frac{1}{9\times10^{-8}}$$

$$= -1.09 - 0.02955 \times \log(1.11\times10^7)$$

= –1.09–0.02955(0.0453+7)
= –1.09–0.208
= –1.298 V

Assertion and Reason Answers:

1. (c) Assertion is correct statement but reason is wrong statement.

Explanation:

Cu^{2+} ions are deposited as Cu.

2. (a) Assertion and reason both are correct statements and reason is correct explanation for assertion.

Case Study Answers:

1. Answer :

i. (b) $E_{cell} > 0; \Delta G < 0$

Explanation:

$$\underset{(1M)}{M} \longrightarrow \underset{(0.05M)}{M^+} + e^-$$

For concentration cell, $E_{cell} = \frac{0.059}{1} \log \frac{0.05}{1}$

$$E_{cell} = \frac{0.059}{1} \log(5 \times 10^{-2})$$

$$E_{cell} = \frac{0.059}{1}[(-2) + \log 5] - 0.059(-2 + 0.698)$$

$$= -0.059(-1.302) = 0.0768$$

$$\Delta G = -nFE_{cell}$$

If E_{cell} is positive, ΔG is negative.

ii. (c) 154mV

Explanation:

$$\frac{E_1}{E_2} = \frac{\log 0.05}{\log 0.0025}$$

$$\frac{E_1}{E_2} = \frac{\log 5\times10^{-2}}{\log 25\times10^{-4}}$$

$$E_1 = 0.0768$$

$$\frac{0.0168}{E_2} = \frac{-1.3}{-2.6} = \frac{1}{2} \text{ or } E_2 = 154\text{mV}$$

iii. (c) > 1

Explanation:

$$K = \text{antilog}\left(\frac{nE^\circ}{0.0591}\right)$$

For feasible cell, E° is positive,
hence from the above equation,
K > 1 for a feasible cell reaction.

iv. (b) 0

v. (a) Concentration ofions in solution.

2. Answer :

i. (b) 1.0

Explanation:

$$n_{NaCl} = \frac{4\times500}{1000} = 2\text{mol}$$

$$\therefore n_{Cl_2} = 1\text{mol}$$

ii. (b) 446g

Explanation:

n_{Na} deposited = 2mol

$\therefore n_{Na-Hg}$ formed = 2 mol

$\therefore$ Mass of amalgam formed = 2 × 223 = 446g

iii. (d) 193000

Explanation:

$$2Na^{+} + \underset{(2F)}{2e^{-}} \rightarrow 2Na$$

Total charge required = 2F = 2 × 96500 = 193000C

iv. (a) 2

v. (a) H_2gas

Important Questions

Multiple Choice questions-

Question 1. If the conductivity and conductance of a solution is same then its cell constant is equal to:

(a) 1

(b) 0

(c) 10

(d) 1000

Question 2.The units of conductivity are:

(a) ohm^{-1}

(b) ohm^{-1} cm^{-1}

(c) ohm^{-2} cm^2 $equiv^{-1}$

(d) ohm^{-1} cm^2

Question 3.The resistance of 0.1 N solution of acetic acid is 250 ohm, when measured in a cell of cell constant 1.15 cm^{-1}. The equivalent conductance (in ohm^{-1} cm^2 $equivalent^{-1}$) of 0.1 N acetic acid is

(a) 18.4

(b) 0.023

(c) 46

(d) 9.2

Question 4.In infinite dilution of aqueous solution of BaCl2, molar conductivity of Ba^{2+} and Cl^- ions are = 127.32 S cm^2/mol and 76.34 S cm^2/mol respectively. What is A°m for BaCl2 at same dilution?

(a) 280 S cm^2 mol^{-1}

(b) 330.98 S cm^2 mol^{-1}

(c) 90.98 S cm^2 mol^{-1}

(d) 203.6 S cm^2 mol^{-1}

Question 5.The specific conductance of 0.1 M NaCl solution is 1.06×10^{-2} ohm^{-1} cm^{-1}. Its molar conductance in ohm^{-1} cm^2 mol^{-1} is

(a) 1.06×10^2

(b)1.06×10^3

(c) 1.06×10^4

(d) 53

Question 6.The limiting molar conductivities A° for NaCl, KBr and KCl are 126, 152 and 150 S cm^2 mol^{-1} respectively. The A° for NaBr is

(a) 278 S cm^2 mol^{-1}

(b) 976 S cm^2 mol^{-1}

(c) 128 S cm^2 mol^{-1}

(d) 302 S cm^2 mol^{-1}

Question 7. λ(ClCH2COONa) = 224 ohm^{-1} cm^2 gm eq^{-1}, λ(NaCl) = 38.2 ohm^{-1} cm^2 gm eq^{-1}. λ(HCl) = 203 ohm^{-1} cm^2 gm eq^{-1}, what is the value of λ(ClCH2COOH)?

(a) 288.5 ohm^{-1} cm^2 gm eq^{-1}

(b) 289.5 ohm^{-1} cm^2 gm eq^{-1}

(c) 388.8 ohm^{-1} cm^2 gm eq^{-1}

(d) 59.5 ohm^{-1} cm^2 gm eq^{-1}

Question 8. The limiting molar conductivities of HCl, CH_3COONa and NaCl are respectively 425, 90 and 125 mho cm² mol^{-1} at 25 °C. The molar conductivity of 0.1 M CH3COOH solution is 7.8 mho cm² mol^{-1} at the same temperature. The degree of dissociation of 0.1 M acetic acid solution at the same temperature is

(a) 0.10

(b) 0.02

(c) 0.15

(d) 0.03

Question 9.The values of limiting ionic conductance of H and HCOO– ions are respectively 347 and 53 S cm² mol^{-1} at 298 K. If the molar conductance of 0.025 M methanoic acid at 298 K is 40 S cm² mol^{-1}, the dissociation constant of methanoic acid at 298 K is

(a) 1×10^{-5}

(b) 2×10^{-5}

(c) $1.5 \times 10\text{-}^{4}$

(d) 2.5×10^{-4}

Question 10.The ionisation constant of a weak electrolyte is 2.5×10^{-5} and molar conductance of its 0.01 M solution is 19.6 S cm² mol^{-1}. The molar conductance at infinite dilution (S cm² mol^{-1}) is

(a) 402

(b) 392

(c) 306

(d) 39.2

Very Short Question:

Question 1. Can you store AgCl solution in Zinc pot?

Question 2. Define the term – standard electrode potential?

Question 3. What is electromotive force of a cell?

Question 4. Can an electrochemical cell act as electrolytic cell? How?

Question 5. Single electrode potential cannot be determined. Why?

Question 6. What is SHE? What is its electrode potential?

Question 7. What does the positive value of standard electrode potential indicate?

Question 8. What is an electrochemical series? How does it predict the feasibility of a certain redox reaction?

Question 9. Give some uses of electrochemical cells?

Question 10. State the factors that affect the value of electrode potential?

Short Questions:

Question 1. What is the cell potential for the cell at $25^0C \quad Cr / Cr^{3+} 10.1\ m] // Fe^{2+} (0.01m) / Fe$

$E^0_{cr+/cr} = -0.74V$. $E^0 Fe^{2+} / Fe = -0.44V$

Question 2. Calculate ΔG^0 for the reaction

$25^0C \ Zn(s) 1 Zn^2 + [0.0004m] 11 cd^2 + (0.2m) 1 cd (s) \ E^0{}_{Zn}{}^{2+}{}_{/Zn} = -0.763V$

$E^0{}_{cd}{}^{2+} / cd = -0.403v,\ F = 96500\ C Mol^{-1}$ R = 8.314 J/K

Question 3. Calculate Equilibrium constant K for the reaction

$298K \ Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+} / aq) + Cu \ E^0{}_{Zn}{}^{2+} 1 Zn = -0.076v; E^0 Cu^2 + / Cu$ +0.34v.

Question 4. For what concentration of $Ag + (aq)$ will the emf of the given cell be zero at 25^0C if the concentration of $CU^{2+}(aq)$ is 0.1 M? $Cu(s) / Cu^{2+}(0.1M) // Ag^+(aq) / Ag(s)$
$E^0 Ag^+ / Ag = +0.80V;\ E^0{}_{Cu}{}^{2+/Cu} = 0.34\ V$

Question 5. Calculate the standard free energy change for the cell- reaction.

$Fe^{2+}(aq) + Ag^+(s) \grave{a} \rightarrow Fe^3 + (aq) + Ag(s)$ How is it related to the equilibrium

constant of the reaction? $E^0{}_{Fe}{}^{3+} / Fe^{2+} = +0.77V$, $E^0{}_{Ag}{}^{+1 Ag} = +0.08V \ F = 96500\ C / mol.$

Question 6. How much charge is required for the following reductions:

(i) 1 mol of to Al.
(ii) 1 mol of to Cu.
(iii) 1 mol of to .

Question 7. How much electricity in terms of Faraday is required to produce

(i) 20.0 g of Ca from molten .
(ii) 40.0 g of Al from molten

Question 8. How much electricity is required in coulomb for the oxidation of

(i) 1 mol of to .
(ii) 1 mol of FeO to

Question 9. A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Question 10. Depict the galvanic cell in which the reaction takes place. Further show:

(i) Which of the electrode is negatively charged?
(ii) The carriers of the current in the cell.
(iii) Individual reaction at each electrode.

Long Questions:

Question 1. Explain construction and working of standard Hydrogen electrode? (b) Write any two differences between amorphous solids and crystalline solids.

Question 2.

The molar conductivity of 0.025 mol L-1 methanoic acid is 46.1 S $cm^2 mol^{-1}$. Calculate its degree of dissociation and dissociation constant. Given $\lambda^\circ H^+$ = 349.6 S $cm^2 mol^{-1}$ and $\lambda^\circ(HCOO-) = 54.6 S\,cm^2 mol$

Question 3.Explain how rusting of iron is envisaged as setting up of an electrochemical cell.

Question 4 Calculate the standard cell potentials of galvanic cells in which the following reactions take place:

Question 4. Write the Nernst equation and emf of the following cells at 298 K:

Question 5. Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

Assertion and Reason Questions:

1. In these questions, a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

a) Assertion and reason both are correct statements and reason is correct explanation for assertion.
b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.
c) Assertion is correct statement but reason is wrong statement.
d) Assertion is wrong statement but reason is correct statement.

Assertion: At the end of electrolysis using platinum electrodes, an aqueous solution of copper sulphate tums colourless.

Reason: Copper in $CuSO_4$ is converted to $Cu(OH)_2$ during the electrolysis.

2. In these questions, a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

a) Assertion and reason both are correct statements and reason is correct explanation for assertion.
b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.
c) Assertion is correct statement but reason is wrong statement.
d) Assertion is wrong statement but reason is correct statement.

Assertion: Zinc displaces copper from copper sulphate solution.

Reason: Eº of zinc is -0.vV and that of copper is +0.34V.

Case Study Questions:

1. The concentration of potassium ions inside a biological cell is at least twenty times higher than the outside. The resulting potential difference across the cell is important in several processes such as transmission of nerve impulses and maintaining the ion balance. A simple model for such a concentration cell involving a metal M is,

$M_{(s)} \mid M^{+}(aq.; 0.05 \text{ molar}) \mid\mid M^{+}(aq; 1 \text{ molar}) \mid M_{(s)}$

The following questions are multiple choice questions. Choose the most appropriate answer:

(i) For the above cell,

a) $E_{cell} < 0; \Delta G > 0$

b) $E_{cell} > 0; \Delta G < 0$
c) $E_{cell} < 0; \Delta G^{\circ} > 0$
d) $E_{cell} > 0; \Delta G^{\circ} < 0$

(ii) If the 0.05 molar solution of M^+ is replaced by a 0.0025 molar M^+ solution, then the magnitude of the cell potential would be:

a) 130mV
b) 185mV
c) 154mV
d) 600mV

(iii)The value of equilibrium constant for a feasible cell reaction is:

a) < 1
b) = 1
c) > 1
d) Zero

(iv)What is the emf of the cell when the cell reaction attains equilibrium?

a) 1
b) 0
c) > 1
d) < 1

(v) The potential of an electrode change with change in:

a) Concentration ofions in solution.
b) Position of electrodes.
c) Voltage of the cell.
d) All of these.

2. All chemical reactions involve interaction of atoms and molecules. A large number of atoms/molecules are present in a few gram of any chemical compound varying with their atomic/ molecular masses. To handle such large number conveniently, the mole concept was introduced. All electrochemical cell reactions are also based on mole concept. For example, a 4.0 molar aqueous solution of NaCl is prepared and 500mL of this solution is electrolysed. This leads to the evolution of chlorine gas at one of the electrode. The amount of products formed can be calculated by using mole concept.

The following questions are multiple choice questions. Choose the most appropriate answer:

(i) The total number of moles of chlorine gas evolved is:

a) 0.5
b) 1.0
c) 1.5
d) 1.9

(ii) If cathode is a Hg electrode, then the maximum weight of amalgam formed from this solution is:

a) 300g
b) 446g
c) 396g
d) 296g

(iii) The total charge (coulomb) required for complete electrolysis is:

a) 186000
b) 24125
c) 48296
d) 193000

(iv) In the electrolysis, the number of moles of electrons involved are:

a) 2
b) 1
c) 3
d) 4

(v) In electrolysis of aqueous NaCl solution when Pt electrode is taken, then which gas is liberated at cathode?

a) H_2gas
b) Cl_2gas
c) O_2gas
d) None of these.

Answers key

MCQ answers:

1. Answer: (a) 1

2. Answer: (b) ohm^{-1} cm^{-1}

3. Answer: (c) 46

4. Answer: (a) 280 S cm^2 mol^{-1}

5. Answer: (a) 1.06×10^2

6. Answer: (c) 128 S cm^2 mol^{-1}

7. Answer: (c) 388.8 ohm^{-1} cm^2 gm eq^{-1}

8. Answer: (b) 0.02

9. Answer: (d) 2.5×10^{-4}

10. Answer: (b) 392

Very Short Answers:

1. No. We can't store AgCl solution in Zinc pot because standard electrode potential of Zinc is less than silver..
2. When the concentration of all the species involved in a half-cell is unity, then the electrode potential is called standard electrode potential.

3. Answer: Electromotive force of a cell is also called the cell potential. It is the difference between the electrode potentials of the cathode and anode.

$$E_{cell} = E_{cathode} - E_{anode}$$

4. Answer: Yes, An electrochemical cell can be converted into electrolytic cell by applying an external opposite potential greater than its own electrical potential.

5. Answer: A single half cell does not exist independently as reduction and oxidation occur simultaneously therefore single electrode potential cannot be measured.

6. Answer: SHE stands for standard Hydrogen electrode. By convention, its electrode potential is taken as 0 (zero).

7. Answer: The positive value of standard electrode potential indicates that the element gets reduced more easily than ions and its reduced form is more stable than Hydrogen gas.

8. The arrangement of metals and ions in increasing order of their electrode potential values is known as electrochemical series.The reduction half reaction for which the reduction potential is

lower than the other will act as anode and one with greater value will act as cathode. Reverse reaction will not occur.

9. Electrochemical cells are used for determining the
   - pH of solutions
   - solubility product and equilibrium constant
   - in potentiometric titrations
10. Factors affecting electrode potential values are –
    - Concentration of electrolyte
    - Temperature.

Short Answers:

1. **Answer**

The cell reaction is

$$2Cr + 3Fe^{2+} 6e^{-} \rightarrow 2Cr^{3+} + 3Fe$$

Nernst Equation –

$$E_{cell} = \left(E^{0}_{Fe^{2+}/Fe} - E^{0}_{Cr^{3+}/Cr} \right) - \frac{0.059}{6} \log \frac{\left[Cr^{3+}\right]^{2}}{\left[Fe^{2+}\right]^{3}}$$

$$= (-0.44v - (-0.74v) - \frac{0.059}{6} \log \frac{(0.10)^{2}}{(0.01)^{3}}$$

$$= 0.3V - \frac{0.059}{6} \log 10^{4}$$

$$= 0.3V - 0.0394V$$

$$= +0.2606\ V$$

2. **Answer:**

The half-cell reactions are

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

Anode:

Cathode: $Cd^{2}+(aq)+2e^{-}\rightarrow Cd(s)$

Nernst Equation

$$E_{cell}=\left(E^{0}{}_{Cathode}-E^{0}{}_{anode}\right)-\frac{0.059}{n}\log\frac{\left[Zn^{2+}\right]}{\left[Cd^{2+}\right]}$$

= (- 0.403 – (-0.763) –

= 0.36V – 0.0798V = 0.4398V

$\Delta G^{0} = -\mathrm{n\,F\,E^{0}{}_{cell}}$

$$=\frac{-2mol\times 96500\ C}{mol\times 0.4398V}$$

= - 8488 J mol^{-1}

3. Answer:

From the reaction, n =2

$E^{0}{}_{cell}=E^{0}cu^{2}+/cu-E^{0}Zn^{2}+/Zn$

= + 0.34v – (-0.76v) = 1.10V

$E^{0}{}_{cell}=\frac{2.303RT}{nF}\log \mathrm{k_c}$

At 298k, $E^{0}{}_{cell}\times\frac{n}{0.059}\log \mathrm{k_c}$

Log k_c = $E^{0}{}_{cell}\times\frac{n}{0.059}$

$1.10\times\frac{2}{0.059}=37.29$

$K_c = Antilog\ 37.29$

$=1.95\ \times 10^{37}$

4. Answer:

$\left[Ag^{+}\right]=5.3\times 10^{-9}M$

5. Answer :

$E^0_{cell} = 0.03V$

6. Answer

(i)

Therefore, Required charge = 3 F

= 289461 C

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

Therefore, Required charge = 2 F

$= 2 \times 96487$ C

= 192974 C

(iii) $MnO_4^- \rightarrow Mn^{2+}$

i.e., $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$

Therefore, Required charge = 5 F

$= 5 \times 96487$ C

= 482435 C

7. Answer:

.(i) According to the question,

Electricity required to produce 40 g of calcium = 2 F

Therefore, electricity required to produce 20 g of calcium =

= 1 F

(ii) According to the question,

$Al^{3+} + 3e^{-1} \rightarrow \underset{27g}{Al}$

Electricity required to produce 27 g of Al = 3 F

Therefore, electricity required to produce 40 g of Al = $\frac{3\times 40}{27}$F

= 4.44 F

8. Answer:

(i) According to the question,

Now, we can write:

$O^2 \rightarrow \frac{1}{2}O_2 + 2e^-$

Electricity required for the oxidation of 1 mol of H_2O to O_2 = 2 F

$= 2 \times 96487$ C

= 192974 C

(ii) According to the question

$Fe^{2+} \rightarrow Fe^{3+} + e^{-1}$

Electricity required for the oxidation of 1 mol of FeO to Fe_2O_3 = 1 F
= 96487 C

9. Answer :

Given,
Current = 5A
Time = 20×60 = 1200 s
Therefore, $\text{Charge} = \text{current} \times \text{time}$

$= 5 \times 1200$
= 6000 C
According to the reaction,
$Ni^{2+}_{(aq)} + 2e^- \rightarrow \underset{58.7g}{Ni_{(s)}}$
Nickel deposited by 2×96487 C = 58.71 g
Therefore, nickel deposited by 6000 C $= \dfrac{58.71 \times 6000}{2 \times 96487} g$
= 1.825 g
Hence, 1.825 g of nickel will be deposited at the cathode.

10.Answer :

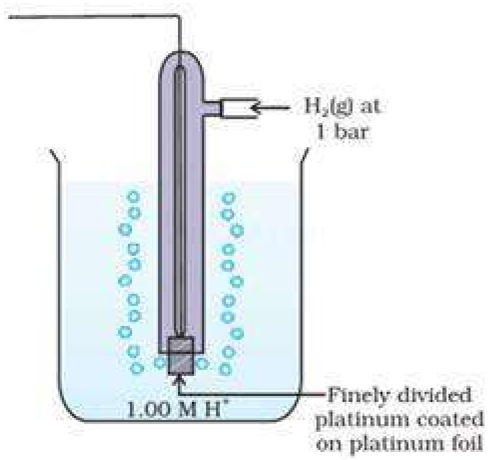
The galvanic cell in which the given reaction takes place is depicted as:
$Zn_{(s)} | Zn^{2+}_{(aq)} || Ag^{+}_{(aq)} | Ag_{(s)}$
(i) Zn electrode (anode) is negatively charged.
(ii) Ions are carriers of current in the cell and in the external circuit, current will flow from silver to zinc.
(iii) The reaction taking place at the anode is given by,
$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$
The reaction taking place at the cathode is given by,
$Ag^{+}_{(aq)} + e^- \rightarrow Ag_{(s)}$

Long Answers:

1. **Answer:**

Construction: SHE consists of a platinum electrode coated with platinum black. The electrode is dipped in an acidic solution and pure Hydrogen gas is bubbled through it. The concentration of both the reduced and oxidized. Forms of Hydrogen is maintained at unity i.e) pressure of gas is 1 bar and concentration of Hydrogen ions in the solution is 1 molar.

Working – The reaction taking place in SHE is At 298 K , the emf of the cell constructed by taking SHE as anode and other half-cell as cathode, gives the reduction potential of the other half cell whereas for a cell constructed by taking SHE as anode gives the oxidation potential of other half cell as conventionally the electrode potential of SHE is zero.



2. Answer:

$\Lambda_m = 46.1 \, S cm^2 mol^{-1}$

$\lambda^\circ(H^+) = 349.6 \, S cm^2 mol^{-1}$

$\lambda^\circ(HCOO^-) = 54.6 \, S cm^2 mol^{-1}$

$\Lambda_m^\circ(HCOOH) = \lambda^\circ(H^+) + \lambda^\circ(HCOO^-)$

= 349.6 + 54.6 = $404.2 \, S cm^2 mol^{-1}$

Now, degree of dissociation:

$$\alpha = \frac{\Lambda_m(HCOOH)}{\Lambda_m^0(HCOOH)}$$

$= \frac{46.1}{404.2}$ = 0.114 (approximately)

Thus, dissociation constant:

$$K = \frac{c\alpha^2}{(1-\alpha)}$$

$$= \frac{(0.025\,\text{mol L}^{-1})(0.114)^2}{(1-0.114)}$$

$$= 3.67\times10^{-4}\,\text{mol L}^{-1}$$

3. **Answer:**

In the process of corrosion, due to the presence of air and moisture, oxidation takes place at a particular spot of an object made of iron. That spot behaves as the anode. The reaction at the anode is given by,

Electrons released at the anodic spot move through the metallic object and go to another spot of the object.

There, in the presence of ions, the electrons reduce oxygen. This spot behaves as the cathode. These H+ ions come either from , which are formed due to the dissolution of carbon dioxide from air into water or from the dissolution of other acidic oxides from the atmosphere in water.

The reaction corresponding at the cathode is given

$$O_{2(g)} + 4H^{+}_{(aq)} + 4e^{-} \rightarrow 2H_2O_{(l)}$$

The overall reaction is: $2Fe_{(s)} + O_{2(g)} + 4H^{+}_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + 2H_2O_{(l)}$
Also, ferrous ions are further oxidized by atmospheric oxygen to ferric ions. These ferric ions combine with moisture, present in the surroundings, to form hydrated ferric oxide $(Fe_2O_3.xH_2O)$ i.e., rust.
Hence, the rusting of iron is envisaged as the setting up of an electrochemical cell.

4. **Answer (i)** $E^\circ_{Cr^{3+}/Cr} = 0.74V$
$E^\circ_{Cd^{2+}/Cd} = 0.40V$

The galvanic cell of the given reaction is depicted as: $Cr_{(s)} \,|\, Cr^{3+}_{(aq)} \,||\, Cd^{2+}_{(aq)} \,|\, Cd_{(s)}$

Now, the standard cell potential is $E^\circ_{cell} = E^\circ_R - E^\circ_L$
= 0.40 – (–0.74)
= +0.34 V
$\Delta_r G^\circ = -nFE^\circ_{cell}$
In the given equation,
n = 6
F = 96487 C mol – 1
E°_{cell} = +0.34 V
Then, $\Delta_r G^\circ = -\ 6\times 96487\text{ C mol}^{-1}\times 0.34\text{ V}$
= –196833.48 CV mol^{-1}

= –196833.48 J mol^{-1}
= –196.83 kJ v
Again, $\Delta_r G^\circ = -RT \ln K$
$\Delta_r G^\circ = -2.303 RT \ln K$
$\log K = -\frac{\Delta_r G}{2.303 RT}$
$= \frac{196.83 \times 10^3}{2.303 \times 8.314 \times 298}$
= 34.496
Therefore, K = antilog (34.496)
$= 3.13 \times 1034$

(ii) $E^\circ_{Fe^{3+}/Fe^{2+}} = 0.77 V$
$E^\circ_{Ag^+/Ag} = 0.80 V$

The galvanic cell of the given reaction is depicted as: $Fe^{2+}_{(aq)} | Fe^{3+}_{(aq)} || Ag^+_{(aq)} | Ag_{(s)}$
Now, the standard cell potential is $E^\circ_{cell} = E^\circ_R - E^\circ_L$
= 0.80 – 0.77
= 0.03 V
Here, n = 1.
Then, $\Delta_r G^\circ = -nFE^\circ_{cell}$
$= -1 \times 96487 \text{ C mol}^{-1} \times 0.03 \text{ V}$
= –2894.61 J mol^{-1}
= –2.89 kJ mol^{-1}
Again, $\Delta_r G^\circ = 2.303 RT \ln K$
$\log K = -\frac{\Delta_r G}{2.303 RT}$
$= \frac{-2894.61}{2.303 \times 8.314 \times 298}$
= 0.5073
Therefore, K = antilog (0.5073)
= 3.2 (approximately)

5. Answer (i) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log \frac{[Mg^{2+}]}{[Cu^{2+}]}$$

$$= \{0.34-(-236)\} - \frac{0.0591}{2}\log\frac{.001}{.0001}$$

$$= 2.7 - \frac{0.0591}{2}\log 10$$

= 2.7 – 0.02955

= 2.67 V (approximately)

(ii) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{n}\log\frac{[Fe^{2+}]}{[H^{+}]^2}$$

$$= \{0-(-0.44)\} - \frac{0.0591}{2}\log\frac{0.001}{1^2}$$

= 0.44–0.02955(–3)

= 0.52865 V

= 0.53 V (approximately)

(iii) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{n}\log\frac{[Sn^{2+}]}{[H^{+}]^2}$$

$$= \{0-(-0.14)\} - \frac{0.0591}{2}\log\frac{0.050}{(0.020)^2}$$

$= 0.14 - 0.0295 \times \log 125$

= 0.14–0.062

= 0.078 V

= 0.08 V (approximately)

(iv) For the given reaction, the Nernst equation can be given as:

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{n}\log\frac{1}{[Br^{-}]^2[H^{+}]^2}$$

$$= (0-1.09) - \frac{0.0591}{2}\log\frac{1}{(0.010)^2(0.030)^2}$$

$$= -1.09 - 0.02955 \times \log\frac{1}{0.00000009}$$

$$= -1.09 - 0.02955 \times \log\frac{1}{9\times10^{-8}}$$

$$= -1.09 - 0.02955 \times \log(1.11\times10^{7})$$

= –1.09–0.02955(0.0453+7)

= –1.09–0.208

= –1.298 V

Assertion and Reason Answers:

1. (c) Assertion is correct statement but reason is wrong statement.

Explanation:

Cu^{2+} ions are deposited as Cu.

2. (a) Assertion and reason both are correct statements and reason is correct explanation for assertion.

Case Study Answers:

1. Answer :

i. (b) $E_{cell} > 0; \Delta G < 0$

Explanation:

$$\underset{(1M)}{M} \longrightarrow \underset{(0.05M)}{M^+} + e^-$$

For concentration cell, $E_{cell} = \frac{0.059}{1} \log \frac{0.05}{1}$

$$E_{cell} = \frac{0.059}{1} \log(5 \times 10^{-2})$$

$$E_{cell} = \frac{0.059}{1}[(-2) + \log 5] - 0.059(-2 + 0.698)$$

$$= -0.059(-1.302) = 0.0768$$

$$\Delta G = -nFE_{cell}$$

If E_{cell} is positive, ΔG is negative.

ii. (c) 154mV

Explanation:

$$\frac{E_1}{E_2} = \frac{\log 0.05}{\log 0.0025}$$

$$\frac{E_1}{E_2} = \frac{\log 5\times10^{-2}}{\log 25\times10^{-4}}$$

$$E_1 = 0.0768$$

$$\frac{0.0168}{E_2} = \frac{-1.3}{-2.6} = \frac{1}{2} \text{ or } E_2 = 154\mathrm{mV}$$

iii. (c) > 1

Explanation:

$$K = \operatorname{antilog}\left(\frac{nE^{\circ}}{0.0591}\right)$$

For feasible cell, E^o is positive,
hence from the above equation,
K > 1 for a feasible cell reaction.

iv. (b) 0

v. (a) Concentration ofions in solution.

2. Answer :

i. (b) 1.0

Explanation:

$$n_{NaCl} = \frac{4\times500}{1000} = 2\mathrm{mol}$$

$$\therefore n_{Cl_2} = 1\mathrm{mol}$$

ii. (b) 446g

Explanation:

n_{Na} deposited = 2mol

$\therefore n_{Na-Hg}$ formed = 2 mol

$\therefore$ Mass of amalgam formed = 2 × 223 = 446g

iii. (d) 193000

Explanation:

$$\underset{}{2Na^+} + \underset{(2F)}{2e^-} \rightarrow 2Na$$

Total charge required = 2F = 2 × 96500 = 193000C

iv. (a) 2

v. (a) H_2gas