Chemistry Worksheet on Chapter 3 Electrochemistry with Answers-Set 1

Q-1: The standard reduction potentials of three metallic cations, A, B, and C, are +0.65, -2.49, and -1.18 V, respectively. The order of the corresponding metal's reducing power is

a) B>C>A
b) A>C>B
c) C>B>A
d) C>A>B

Answer: a) B>C>A

Explanation: Reducing power refers to a metal's ability to cause reduction while also undergoing oxidation itself.

Reducing power is inversely proportional to the reduction potential value of the metallic cations.More the reduction potential less is the reducing power.

The correct order is B>C>A

Q-2: The reaction

$$\frac{1}{2} H_2(g) + AgCl(s) \rightarrow H^+(aq) + Cl^-(aq) + Ag(s)$$

occurs in the galvanic cell

a) Ag/AgCl(s)|KCl(soln)|$AgNO_3$|Ag
b) Pt/H_2(g)|KCl(soln)|AgCl(s)\Ag
c) Pt|H_2(g)|HCl(soln)|AgCl(s)|Ag
d) Pt|H_2(g)|HCl(soln)|$AgNO_3$(soln)|Ag

Answer: c) Pt|H_2(g)|HCl(soln)|AgCl(s)|Ag

Q-3: Pure water does not conduct electricity because

a) The melting point is low.
b) Is easily broken down
c) Is almost completely unionised
d) Is neutral

Answer: c) Is almost completely unionised

Explanation: It is essential to understand that electrical conduction is the movement of electrically charged particles through a transmission medium, which implies that conduction necessitates the presence of free ions. Pure water has a very low ion concentration and is almost unionised, so it does not conduct electricity.

Q-4: Calculate the charge required to oxidise one mole of Mn_3O_4 into MnO_4^{2-} in the presence of an alkaline medium.

Answer: Charge, $Q = n \times n_f \times F$

Where n is the number of moles
n_f is the change in electrons that occurs during oxidation/reduction
F is the Faraday's constant

In the reaction, $Mn_3O_4 \rightarrow 3MnO_4^{2-}$

Mn is in 8/3 oxidation state in Mn_3O_4, and in MnO_4^{2-} it is in +6 oxidation state.
$n_{f.} = [6-(8/3)] \times 3 = 10$

Substituting values in the expression for charge, Q.
$Q= 1\times 10 \times 96500 = 965000$ C

Q-5: The amount of substance liberated at an electrode is not directly proportional to:

a) Time
b) Current
c) Electrochemical equivalent
d) Conductivity

Answer: d) Conductivity

Explanation: According to Faraday's first law.

$w= z\times I\times t$

Where w is the amount of substance liberated
Z is the electrochemical equivalent
I is the current
t is the time

We can clearly see that the amount of substance liberated is directly proportional to current, time, and electrochemical equivalent and not conductivity.

Q-6: In a cell, $Zn(s)|Zn^{2+}||\ H^+|H_2(Pt)$; adding H_2SO_4 to the cathode compartment will result in

a) Decrease E_{cell}
b) Increase E_{cell}
c) Adjust the equilibrium to the left
d) Adjust the equilibrium to the right

Answer: b) and d)

Explanation: For the cell: $Zn(s)|Zn^{2+}||\ H^+|H_2(Pt)$

Electrode	Reaction
Anode	$Zn \rightarrow Zn^{2+} + 2e^-$
Cathode	$2H^+ + 2e^- \rightarrow H_2$

Overall reaction: $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$

Applying Nernst equation,

$$E_{cell} = E^o_{cell} - \frac{0.0591}{n} log \frac{P_{H_2}[Zn^{2+}]}{[H^+]^2}$$

The Nernst equation clearly shows that increasing H_2SO_4 concentration (which means increasing $[H^+]$) increases E_{cell}.

According to Le Chatelier's Principles, increasing the reactant concentration causes equilibrium to shift to the right. We can see from the overall reaction that $[H^+]$ is present on the reactant side, indicating that equilibrium will shift to the right.

Q-7: The E^0 values for the two reduction electrode processes are as follows:

a) $Ce^{4+}/Ce^{3+} = 1.61$ V
b) $Co^{2+}/Co = -0.28$V

Calculate the cell potential and ΔG^o for the cell reaction.

Answer: For a cell reaction to be spontaneous, its cell potential (E^0_{cell}) must be positive. It is only possible if cobalt is oxidised on the anode and Ce^{4+} is reduced on the cathode.

With this, the correct cell representation will be : $Co|Co^{2+}||Ce^{4+}|Ce^{3+}$

We are given reduction potentials for both the electrodes.

$E^0_{Ce^{4+}/Ce^{3+}}$ = 1.61 V

$E^0_{Co/Co^{2+}} = -E^0_{Co^{2+}/Co}$ = 0.28V

$E^0_{cell} = E^0_{Ce^{4+}/Ce^{3+}} + E^0_{Co/Co^{2+}}$ = 1.61 V + 0.28V = 1.89 V

The half cell reactions are:

$Ce^{4+} + e^- \rightarrow Ce^{3+}$] ×2

$Co \rightarrow Co^{2+} + 2e^-$

So the overall reaction is

$2Ce^{4+} + Co \rightarrow 2Ce^{3+} + Co^{2+}$

and the number of electrons involved, n= 2.

$\Delta G^o = -nFE^0_{cell} = -2 \times 96500 \times 1.89 = -364770$ J/mol

Q-8: Define the following:

a) Strong Electrolyte
b) Kohlraushís law
c) Molar conductivity

Answer:

a) A strong electrolyte is a substance that is completely ionised in solution. NaCl is an example.

b) It states that for a given salt, the molar conductivity at infinite dilution can be expressed as the sum of the individual contributions from the electrolyte ions.

c) Molar conductivity is the conductance of all the ions produced by one mole of electrolyte when the electrodes are separated by a unit distance and have a sufficient area of cross-section to hold the electrolyte.

Q-9: The following are the standard reduction potentials for single electrodes at 298 K:

Electrode	Electrode potential (Volt)
Mg^{2+}/Mg	-2.34
Pb^{2+}/Pb	-0.1262
Cr^{2+}/Cr	-0.913

From this we can infer that

a) Pb can reduce both Mg^{2+} and Cr^{2+}
b) Cr can reduce both Mg^{2+} and Pb^{2+}
c) Mg can reduce both Pb^{2+} and Cr^{2+}
d) Mg can reduce both Pb^{2+} but not Cr^{2+}

Answer: c) Mg can reduce both Pb^{2+} and Cr^{2+}

Explanation: Metals with lower SRP can displace the metal with higher SRP from their solutions..Mg with lower SRP(-2.34) can reduce both Pb^{2+}(-0.1262) and Cr^{2+}(-0.913).

Q-10: Given that $E^0(Fe^{3+}, Fe)$= -0.4 V and $E^0(Fe^{2+}, Fe)$=-0.44V, the value of $E^0(Fe^{3+},Fe^{2+})$ is

a) 0.76 V
b) -0.40 V
c) -0.76 V
d) 0.40 V

Answer: a) 0.76 V

Explanation:

$Fe^{3+} + 3e^- \rightarrow Fe$; $E^0 = -0.4$ V ………………1)

$\Delta G^o_1 = -3 \times F \times -0.4 = +0.12F$

$Fe^{2+} + 2e^- \rightarrow Fe$; $E^0 = -0.44$ V
On inverting we get
$Fe \rightarrow Fe^{2+} + 2e^-$; $E^0 = 0.44$ V …………………2)
$\Delta G^o_2 = -2 \times F \times 0.44 = -0.88$ F

On adding 1 and 2, we get

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

$\Delta G^o = \Delta G^o{}_1 + \Delta G^o{}_2 = +0.12F + (-0.88\ F) = -0.76\ F$

$\Delta G^o = -nFE^0{}_{cell} = -1 \times F \times E^0{}_{cell}$
$-0.76 = -1 \times F \times E^0{}_{cell}$

$E^0{}_{cell}{}^{=} = 0.76V$

Q-11: The Nernst equation for the reaction, $A^{2+} + 2e^{-} \rightarrow B$ in terms of the free energy change is

a) $\Delta G = \Delta G^o + 2.303RT \log([B]/[A])$
b) $\Delta G = \Delta G^o - 2.303RT \log([B]/[A])$
c) $-\Delta G = -\Delta G^o + 2.303RT \log([B]/[A])$
d) $\Delta G = -\Delta G^o + 2.303RT \log([B]/[A])$

Answer: a) $\Delta G = \Delta G^o + 2.303RT \log([B]/[A])$

Q-12: The standard Gibbs free energy change for the reaction shown below is -2.7kJ/mol

$Sn(s) + Pb^{2+} \rightarrow Sn^{2+} + Pb(s)$

Given that $E^0(Pb^{2+}/Pb)$ is -0.126 V the value of $E^0(Sn^{2+}/Sn)$ in V is ____________(upto two decimal places)

Answer: Given: $\Delta G =$ -2.7kJ/mol = 2700J/mol

$E^0(Pb^{2+}/Pb)$ is -0.126 V
$E^0(Sn^{2+}/Sn) = ?$

According to the equation, Sn is oxidised to Sn^{2+} by losing two electrons, whereas Pb^{2+} is reduced to Pb by gaining two electrons.

$\Delta G = -nFE^0{}_{cell}$
$-2700J = 2 \times 96500C \times E^0{}_{cell}$

$E^0{}_{cell} = 0.01398\ V$

$E^0{}_{cell} = E^0(Pb^{2+}/Pb) - E^0(Sn^{2+}/Sn)$

$0.01398 = -0.126\ V - E^0(Sn^{2+}/Sn)$

$E^0(Sn^{2+}/Sn)$= -0.13V

Q-13: A column of 0.05M NaOH solution with a diameter of 1 cm and a length of 50 cm has an electrical resistance of 5.55 ×10^3 ohm. Determine the resistivity, conductivity, and molar conductivity.

Answer: Diameter= 1 cm, radius = 0.5 cm

Area =πr^2 = 3.14 × $(0.5)^2$ = 0.785 cm^2

Resistivity,ρ= R×A/l
ρ= (5.55 ×10^3×0.785) /50

On solving, ρ**= 87.135 ohm cm**

Conductivity,κ= 1/ρ

κ**= 0.01148 $ohm^{-1}cm^{-1}$**

Molar conductivity,Λ_m^c = κ×1000/M
Where M is the concentration of NaOH solution

Λ_m^c = (0.01148×1000/0.05) = 229.6 $ohm^{-1}cm^2$/mol

Q-14: Why is it not possible to measure the potential of a single electrode?

Answer: Because the half cell containing single electrode cannot exist independently, as charge cannot flow on its own in a single electrode.

Q-15: Explain the distinction between primary and secondary batteries.

Answer:

Primary Batteries	**Secondary Batteries**
The reaction occurs only once in primary batteries, and after prolonged use, the battery dies and cannot be reused.	After use, a secondary cell can be recharged by passing current through it in the opposite direction, allowing it to be reused. A good secondary cell can withstand numerous discharging and charging cycles.
Example: Leclanche cell	Example: Lead Storage battery

Q-16: For the following reaction,
$2MnO_4^- + 5H_2C_2O_4 + 6H^+ \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$
$E^0(MnO_4^-/Mn^{2+})$= +1.51 V and $E^0(CO_2/H_2C_2O_4)$ = -0.49 V. Calculate the equilibrium constant at 298 K.

Answer: $E^0_{cell} = E^0_{(Cathode)} - E^0_{(anode)} = 1.51-(-0.49) = 2V$

The half cell reaction are:

$MnO_4^- + 5e^- \rightarrow Mn^{2+}] \times 2$
$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-] \times 5$

So the balanced equation is
$2MnO_4^- + 5H_2C_2O_4 + 6H^+ \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$
and the number of electrons involved, n= 10.

At 298 K,

$E^0_{cell} = (0.0591/n)\log(K_{eq})$
$2 = (0.0591/10)\log(K_{eq})$
$\log(K_{eq}) = 338$
$K_{eq} = 10^{338}$

Q-17: The electrical conductivity of a metal

a) Increases with increasing temperature
b) Decreases with increasing temperature
c) Is independent of temperature
d) Shows oscillatory behaviour with temperature

Answer: b) Decreases with increasing temperature

Explanation: In the case of metals, the current carrying particles are electrons,that is, a systematic flow of electrons in one direction will be responsible for the conduction of electricity. When temperature increases, some of the electrons get excited and this causes them to move in a not so orderly manner.

Hence the electrons become less efficient as the charge carries and result in decreased conductivity.

Q-18: The mean ionic activity coefficient of 0.001 molal $ZnSO_4$(aq) at 298 K according to the Debye-Huckel limiting law is(Debye Huckel constant is 0.509 molal$^{-1/2}$)

Answer: Mean ionic activity,I = ½ $(m_1z_1^2 + m_2z_2^2)$

Where m_1 and m_2 are the molality of Zn^{2+} and SO_4^{2-} respectively

z_1= charge on Zn^{2+}
z_2 = charge on SO_4^{2-}

Mean ionic activity,I =½(0.001(2)^2+ 0.001(-2)^2)= 0.004

We know that, $\log \gamma\pm = -0.509|z_1z_2|\ I^{1/2}$
$\log \gamma\pm = -0.509|(+2)(-2)|0.004^{1/2}$
$\gamma\pm = 0.743$

Where $\gamma\pm$ is mean ionic coefficient

Q-19: When an Al-NaCl solution is electrolyzed, how does the pH change?

Answer: When an Al-NaCl solution is electrolyzed, H_2 is released at the cathode, Cl_2 is released at the anode, and NaOH is formed in the solution. As a result of the formation of the base(NaOH), the pH of the solution rises.

Q-20: The H_2 -O_2 fuel cell was used in the Apollo space programmes.

(a) Explain why fuel cells are preferred in space missions.
(b) Discuss the values that influenced the decision to use fuel cells.
(c) Can fuel cells be used in automobiles?
(d) How can we increase efficiency of fuel cells ?

Answer:
a) Fuel cells use the reaction of hydrogen with oxygen to produce water. The cell was used to provide electrical power during the Apollo space programme.The water vapours produced by the reaction were condensed and added to the astronauts' drinking water supply.

b) Better knowledge and understanding of fuel cells, including their benefits and efficiency, influenced the decision to use them.

c) Yes, These have been used experimentally in automobiles. Fuel cells are pollution-free, and a variety of fuel cells have been manufactured and tested in anticipation of their future importance.

d) For increasing the efficiency of fuel cells, tremendous progress has been made in the development of new electrode materials, better catalysts, and electrolytes.