Nernst Equation Chemistry Questions with Solutions

Q1: Calculate the equilibrium constant for the reaction Fe + $CuSO_4 \rightleftharpoons FeSO_4$ + Cu at 25°C.
(Given E°(OP/Fe) = 0.5 V°, E°(OP/Cu) = −0.4 V)
a) 3.46 × 10^{30}
b) 3.46 × 10^{26}
c) 3.22 × 10^{30}
d) 3.22 × 10^{26}

Answer: c) 3.22 × 10^{30}

Explanation: The cell reaction shows the oxidation of Fe and reduction of Cu^{2+}

Therefore for the reaction, Fe + $CuSO_4 \rightleftharpoons FeSO_4$ + Cu

$$E^{\circ}_{(cell)} = E^{\circ}_{(OP/Fe)} + E^{\circ}_{(RP/Cu)}$$

⇒ $E^{\circ}_{(cell)} = 0.5 + 0.4 = 0.9V$

We have, $E^{\circ} = \frac{0.059}{2} log_{10} K_c$

∴ $0.9 = \frac{0.059}{2} log_{10} K_c$

∴ K_c = 3.22 × 10^{30}

Q2: The e.m.f and the standard e.m.f of a cell in the following reaction is 5V and 5.06V at room temperature, $Ni_{(s)} + 2Ag^{+}(n) \rightarrow Ni^{2+}_{(0.02M)} + 2Ag_{(s)}$. What is the concentration of Ag^{+} ions?
a) 0.0125 M
b) 0.0174 M
c) 0.0625 M
d) 0.0314 M

Answer: b) 0.0174 M

Explanation: Given, Temperature T = 298K

Concentration of Ni^{2+} = (0.02M)

$$E_{(cell)} = E^{\circ}_{(cell)} - \frac{0.059}{n} log_{10}(Anode/Cathode)$$

$\Rightarrow$ $5 = 5.06 - \frac{0.059}{2} log_{10}(0.02/[Ag^{+}]^{2})$

$\therefore [Ag^{+}]^{2} = 0.0174M$

Q3: What is the pH of HCl solution when the hydrogen gas electrode shows a potential of −0.22V at standard temperature and pressure?
a) 2.17
b) 2.98
c) 3.14
d) 3.73

Answer: d) 3.73

Explanation: Given, potential of hydrogen gas electrode = −0.22 V

Electrode reaction: $H^{+} + e^{-} \rightarrow 0.5\ H_2$

Applying Nernst equation,

$$E_{(H+/H_2)} = E^{\circ}_{(H+/H_2)} - 0.059 log(1/[H^{+}])$$

$E^{\circ}_{(H+/H_2)}$ = 0 for hydrogen gas electrode

−0.22 = 0.059 log H^{+}

−0.22 = −0.059pH

∴ pH = 3.73.

Q4: Calculate the e.m.f. of the half-cell given below.
Pt, H_2 | HCl at 1-atmosphere pressure and 0.1 M. Given, $E^{\circ}_{(OP)}$ = 2 V.
a) 4 V
b) 5.4 V
c) 3.4 V
d) 5.6 V

Answer: b) 5.4 V

Explanation: Given, $E^\circ_{(OP)}$ = 2V,

$H_2 \rightarrow 2H^+ + 2e$

$$E_{(OP)} = E^\circ_{(OP)} - \frac{0.059}{2} log_{10}([H^+]^2/P(H_2))$$

$$E_{(OP)} = 2 - \frac{0.059}{2} log_{10}(0.02^2/1) = 5.4V$$

Q5: Give the applications of Nernst Equation.

Answer:

The Nernst equation can be used to calculate:

- At any conditions, single electrode reduction or oxidation potential.
- Potentials of standard electrodes.
- The relative ability as a reductive or oxidative agent is compared.
- Determining the possibility of combining single electrodes to generate electric potential.
- Emf of an electrochemical cell.
- Unknown ionic concentrations.
- The Nernst equation can be used to determine the pH of solutions and the solubility of sparingly soluble salts.

Q6: From the following standard potentials, arrange the metals in the order of their increasing reducing power.

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$: E° = −0.76 V
$Ca^{2+}(aq) + 2e^- \rightarrow Ca(s)$: E° = −2.87 V
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$: E° = −2.36 V
$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$: E° = −0.25 V
$Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$: E° = +0.25 V

Answer:

The ability of a metal to give up electrons, i.e. lower standard potentials, enhances its reducing power. The increasing order of reducing power of metals is obtained by arranging the reduction potentials in decreasing order.

Increasing order of reduction potentials is Ni (−0.25V) < Zn (−0.76V) < Mg(−2.36V) < Ca(−2.87).

Q7: Represent the cell in which the following reaction takes place:

$$Mg_{(s)} + 2Ag^{+}_{(0.0001M)} \rightarrow Mg^{2+}_{(0.130M)} + 2Ag_{(s)}$$

Calculate its E_{cell} if E°_{cell} = 3.17 V.

Answer:

The half cell reactions can be written as:

$$Anode : Mg_{(s)} \rightarrow Mg^{2+}_{(0.130M)} + 2e^{-}$$

$$Cathode : 2Ag^{+}_{(0.0001M)} + 2e^{-} \rightarrow 2Ag_{(s)}$$

- The cell can be represented as:

$$Mg_{(s)}/Mg^{2+}_{(0.130M)} \parallel Ag^{+}_{(0.0001M)}/Ag_{(s)}$$

- Calculation of E_{cell} :

Number of electrons exchanges, n = 2

$$ReactionQuotient, Q = \frac{[Mg^{2+}]}{[Ag^{+}]^{2}}$$

$$Q = \frac{0.130}{(1x10^{-4})^{2}} = 1.3x10^{7}$$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} logQ$$

$$E_{cell} = 3.17 - \frac{0.0591}{2} log(1.3x10^{7}) = 2.96V$$

Q8: What is Gibbs energy in simple terms?

Answer:

Gibbs free energy, also known as Gibbs function, Gibbs energy, or free enthalpy, is a term used to measure the maximum amount of work done in a thermodynamic system when temperature and pressure remain constant. The sign 'G' stands for Gibbs free energy.

Q9: Which equation do you use to calculate the membrane potential?

Answer:

The Nernst equation can be used to calculate membrane potential (E). The resting membrane potential, or baseline state, of a cell, is calculated using this equation.

Q10: What are the limitations of the Nernst Equation.

Answer:

1. Since active coefficients in dilute solutions are close to unity, Nernst Equation can be given solely in terms of concentrations. However, at greater concentrations, the real ion activity must be used. This complicates the use of the Nernst Equation because estimating non-ideal ion activity usually requires experimental data.

2. The Nernst Equation is likely to apply when there is no current flow through the electrode. When current flows, the activity of ions at the electrode surface changes, and the measured potential is affected by extra over potential and resistive loss factors.

3. At extremely low concentrations of the potential-determining ions, the potential was discovered using Nernst Equation approaches towards $\pm\infty$. This has no practical relevance since, under such conditions, the exchange current density drops to a very low level, and other factors take over control of the system's electrochemical behaviour.

Q11: What is the difference between the Goldman and Nernst equations?

Answer:

The primary difference between the Nernst equation and the Goldman equation is that the former represents the relationship between reduction potential and standard electrode potential, whilst the latter is a derivative of the Nernst equation and represents the reversal potential across a cell membrane.

	Nernst Equation	**Goldman Equation**
DEFINITION	It is a mathematical expression that gives the relationship between reduction potential and the	It is useful in determining the reverse potential across a cell membrane in a cell

	standard reduction potential of an electrochemical cell.	membrane physiology.
NATURE	Applied for electrochemical cells.	Applied for biological cells.
CONSIDERATIONS	Include the reduction potential, standard reduction potential, temperature and activities of chemical species.	Takes the uneven distribution of ions across the cell membrane and differences in membrane permeability into account.

Q12: How does the Nernst equation calculate equilibrium potential?

Answer:

Based on the charge on the ion (i.e., its valence) and the concentration gradient across the membrane, the Nernst equation derives the equilibrium potential (also known as the Nernst potential) for an ion.

Q13: What is the significance of the Nernst equation?

Answer:

Under non-standard conditions, the Nernst Equation can be used to determine cell potential. It links the observed cell potential to the reaction quotient, allowing for the correct equilibrium constant determination (including solubility constants).

Q14: What is the effect of catalyst on:
(i) Gibbs energy (ΔG) and
(ii) activation energy of a reaction?

Answer:

(i) The catalyst will have no influence on Gibb's energy.
(ii) By lowering the activation energy of a process, the catalyst gives an alternate pathway.

Q15: The chemistry of corrosion of iron is essentially an electrochemical phenomenon. Explain the reactions occurring during the corrosion of iron in the atmosphere.

Answer:

The electrochemical theory is used to explain the corrosion mechanism. We refer to the creation of small electrochemical cells on the surface of iron by using the example of rusting.

The redox reaction involves:

At anode : $Fe(S) \rightarrow Fe^{2+}(aq) + 2e^{-}$
At cathode : $H_2O + CO_2 \rightleftharpoons H_2CO_3$ (Carbonic acid)

$H_2CO_3 \rightleftharpoons 2H^+ + CO_2^{2-}$
$H_2O \rightleftharpoons H^+ + OH^-$
$H^+ + e^- \rightarrow H$
$4H + O_2 \rightarrow 2H_2O$

Then net resultant Redox reaction is
$2Fe(s) + O_2\,(g) + 4H^+ \rightarrow 2Fe^{2+} + 2H_2O$

Practise Questions on Nernst Equation

Q1: What is the number of electrons transferred in an equation if the Nernst equation is $E_{(cell)} = E°_{(cell)} - 9.83 \times 10^{-3} \times \log_{10}$ (Anode / Cathode)?
a) 2
b) 6
c) 4
d) 1

Answer: b) 6

Explanation: Nernst equation = $E^{\circ}_{(cell)} - \frac{0.059}{n} log_{10}(Anode/Cathode)$

On comparing both the formulae, $\frac{0.059}{n} = 9.83x10^{-3}$

∴ n = 6

Q2: Find the number of electrons transferred in the equation $Cu_{(g)} + 2Ag^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$.
a) 4
b) 3
c) 2
d) 1

Answer: c) 2

Explanation: $2Ag^+_{(aq)} + 2e^- \rightarrow 2Ag_{(s)}$

From the equation it is evident that $2Ag^+$ takes 2 electrons from Cu and neutralizes to form 2Ag.

Q3: What is the Nernst equation simplified?

Answer:

The Gibbs free energy is used to derive the Nernst equation. Using the definitions of ΔG = -nFE and ΔG° = -nFE°, we can rewrite this equation. To make things easier, we divide each side by -nF to get the Nernst equation as we know it.

Q4: What is Z in the Nernst equation?

Answer:

The Nernst equation is used to calculate a cell's potential. The number of moles of electrons exchanged in the cell reaction is denoted by n or z (from the German word "Zahl").

Q5: The standard electrode potential for Daniell cell is 1.1V. Calculate the standard Gibbs Energy for the reaction:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

Answer:

Number of electrons exchanged, n = 2

Standard Gibbs Energy of the reaction:

$$\Delta G^0 = -nFE^0_{cell}$$

$$\Delta G^0 = -2x96500x1.1$$

$$\Delta G^0 = -212300J = 212.3KJ$$