ELECTROCHEMISTRY

01/03/2024.

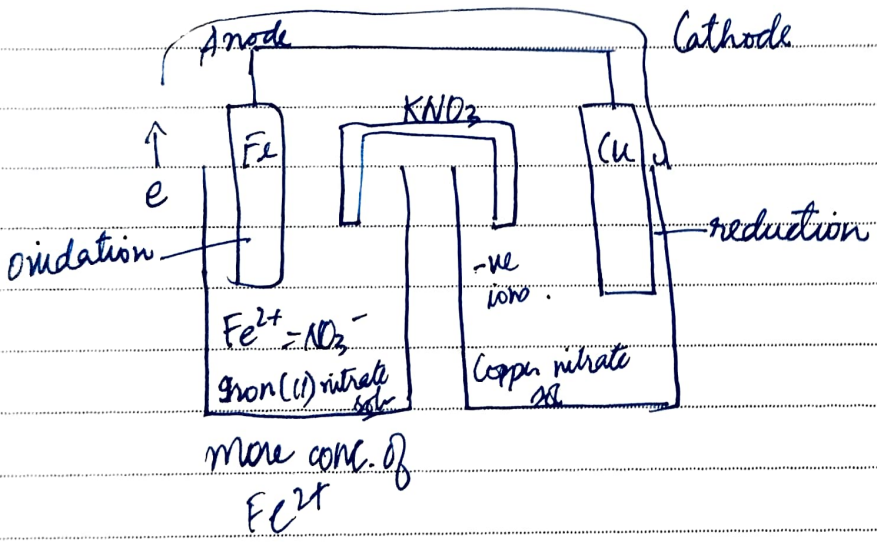
- Electrochemistry is the branch of chemistry which dealts with change in chemical energy into electrical and vice versa.

Redox reaction

Reduction
Gain of e^-
$M^{n+}(aq) + ne^- \longrightarrow M(s)$
eg: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

Oxidation
Loss of e^-
$M(s) \longrightarrow M^{n+}(aq) + ne^-$
eg. $Zn \longrightarrow Zn^{2+} + 2e^-$

Electrochemical cells or Galvanic cells or Daniel cell or Voltaic cell.



L – left side
O – oxidation
A – Anode
N – Negative charge

Functions of salt bridge:

- It completes inner circuit
- It maintains electrical neutrality.

Q. Why in salt bridge KCl & KNO_3 is used as electrolyte.

Ans They are having equal ionic mobility
$K^+ = NO_3^- = Cl^-$

Electrode Potential. in a half cell.
Tendency of an electrode to gain or lose e^-.

i) Oxidation Potential:

Tendency of an electrode to lose e^-. its

eg. $M^0(s) \longrightarrow Mn^+(aq) + ne^-$

$Zn \longrightarrow Zn^{2+} + 2e^- \; 0.76\,V$

ii) Reduction Potential

Tendency of an electrode to gain e^-. its

$Mn^+(aq) + ne \longrightarrow M$

$Cu^+ + 2e^- \longrightarrow (Cu \quad +0.34\,V.$

$$E^\circ cell = E^\circ_C - E^\circ_A$$

$$= E^\circ_{\text{reduction potential}} - E^\circ_{\text{oxidation potential}}$$

$Zn \longrightarrow Zn^{2+} + 2e^-$

$Zn^{2+} + 2e^- \longrightarrow Cu$

Note :- -ve sign in formula automatically change given reduction potential value into oxidation potential.

Q. Predict if the reaction between the following is feasible?

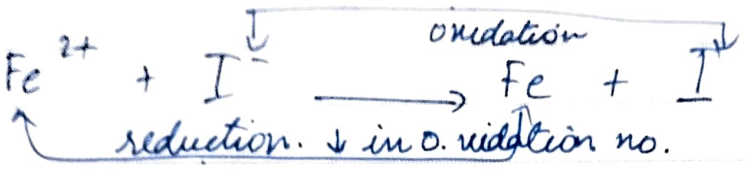
i) $Fe^{2+}(aq)$ and $I^-(aq)$

$E^\circ(Fe^{2+}/Fe) = -0.44\,V.$

$E^\circ(I_2/I^- = 0.54\,V.$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

red – oxidation ↑↑ in oxidation no.



$\therefore -0.44 - 0.54$

$= -0.98\,V.$

E°_{cell} should be +ve for feasible process

$(-ve) \leftarrow \Delta G = -nFE^\circ$

Electrolytic conduction

Conductance

metallic conductor	ionic conductor
① e^- are responsible	① ions are responsible
② Solid state & molten	② Solid state × molten state aqueous state
③ ↓↓ with ↑↑ in temp as ↑↑ in temp cause resistance for flow of e^- as metal atoms vibrates	③ ↑↑ with ↑↑ in temp as ↑↑ in temp ↑↑ rate of dissociation of salt

Electrolytic conductance

Ohm's law

$R = \rho \frac{l}{A}$ where R is resistance
ρ is resistivity
l is length of cell
a is area of cross section

$\frac{1}{\rho} = \frac{1}{R} \frac{l}{A}$

$\frac{1}{\rho} = k$ = specific conductance (k) → unit is $\Omega^{-1} cm^{-1}$ or $S cm^{-1}$

$G = \frac{1}{R}$ = conductance

$k = G \frac{l}{A} = \frac{1}{R} \frac{l}{A}$

$\frac{l}{A}$ → cell constant → unit is cm^{-1}

Molar conductance (λ_m)

$\lambda_m = k \times \frac{1000}{M}$ → molarity

$= k \frac{\Omega^{-1} cm^{-1}}{\text{molarity} \rightarrow mol\, L^{-1}}$ $\quad 1\, L = 1000\, cm^3$

$= \frac{\Omega^{-1} cm^{-1}\, 1L}{mol}$

$= \frac{\Omega^{-1} cm^{-1}\, 1000\, cm^3}{mol}$

$= \Omega^{-1} cm^{-2} mol^{-1}$

8 The electrical resistance of a column of 0.05 mol L^{-1} NaOH solution of diameter 1 cm and length 50 cm is 5.55×10^3 ohm. Calculate its resistivity, conductivity and molar conductivity.

Ans ① Here given $l = 50$ cm

$a = \pi r^2$

$= 3.14 \times (0.5)^2$

$= 3.14 \times 0.25$

$= 0.785\ cm^2$

$\frac{d}{2} = r$

$\frac{1}{2} = 0.5$ cm

$R = \rho \frac{l}{a}$

$\rho = R \times \frac{a}{l}$

$= 5.55 \times 10^3 \times \frac{0.785}{50}$

$= 87.18\ \Omega\ cm$

$\frac{l}{a} = \frac{50}{0.785}$

$= 63.7\ cm^{-1}$

② $\kappa = \frac{1}{\rho} = \frac{1}{87.18} = \frac{1000}{8718} \times 10^{-2}$

$= 1.14 \times 10^{-2}\ S cm^{-1}$

③ $\Lambda_m = \frac{\kappa \times 1000}{M} = \frac{1.14 \times 10^{-2} \times 1000}{0.05 \times 100}$

$= \frac{114}{5} \times \frac{200}{1000}$

$= 228.\ cm^2\ \Omega^{-1}\ mol^{-1}$

Resistance of a conductivity cell filled with 0.1 mol L^{-1} KCl solution is 100 Ω. If the resistance of the same cell when filled with 0.02 mol L^{-1} KCl solution is 520 Ω, calculate the conductivity and molar conductivity of 0.02 mol L^{-1} KCl solution. The conductivity of 0.1 mol L^{-1} KCl solution is 1.29 S/m.

Ans. 0.1 mol L^{-1}

$R = 100\ \Omega$

$R = \rho \frac{l}{a}$

$\frac{1}{\rho} \times R = \frac{l}{a}$

$k \times R = \frac{l}{a}$

$1.29 \times 100 = \frac{l}{a}$

$129 = \frac{l}{a}$

$k = \frac{1}{R} \frac{l}{a}$

$= \frac{1}{520} \times 129$

0.02 mol L^{-1}

$R = 520\ \Omega$

$\Lambda_m = ?$

$k = ?$

$\Lambda_m = \frac{k \times 1000}{M}$

$= \frac{1}{520} \times \frac{129 \times 1000}{0.02}$

$k = \frac{1290}{520} \times 10^{-1}$

$= 0.248$

$= \frac{0.248 \times 1000}{0.02}$

$= 12400$

Effect of dilution on molar conductance

Strong electrolyte

On dilution interionic attraction ↓↓ as a result ionic mobility ↑↑.

Weak electrolyte

molar conductance ↑↑ on dilution as rate of dissociation ↑↑. ions can rejoin with that speed since rate of backward reaction increase.

07/05/2024

Oxidation $Zn \rightarrow Zn^{2+} + 2e^-$

Reduction $Cu^{2+} + 2e^- \rightarrow Cu$

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

reducing agent — oxidising agent.

Factors involving in reducing agent nature.

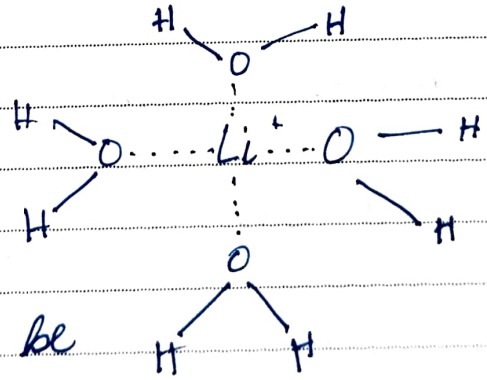
Reducing agent depends on

i) Metal (s) $\xrightarrow{\text{sublimation enthalpy}}$ M (g) $\xrightarrow{\text{ionization enthalpy}}$ $M^+ (g) + e^-$ $\xrightarrow{\text{hydration enthalpy}}$ $M^+ (aq)$

> Sublimation enthalpy
Amount of energy required to convert 1 mol of solid into gaseous phase. It is an endothermic process. [Bond breaking happens]

> Ionization enthalpy
Amount of energy required to remove e^- from outermost shell of an isolated gaseous atom. endothermic

> Hydration enthalpy
Amount of energy released when an ion is surrounded by water molecule. It is an exothermic process.



Bond form → exo
Bond break → endo

Smaller the ion, more will be hydration enthalpy.

$Li^+ > Cs^+$

Ion dipole interaction

Why Lithium is strongest reducing agent? than Cs.

Ans ~~As its hydration~~ Lithium being smallest ion, its hydration enthalpy ~~will be~~ highest which can compensate high sublimation enthalpy as well as ionization enthalpy.

Ans Lithium being smallest ion, its hydration (exothermic) enthalpy is too high which can compensate high sublimation (endothermic) enthalpy as well as ionization (endothermic) enthalpy

Oxidising agent depends on

$$X_2 \xrightarrow{\text{bond dissociation enthalpy}} 2X$$

i) Bond dissociation enthalpy
> Process is endothermic.
> Amount of energy required to break bond in a molecule

$I_2 < F_2 < Br_2 < Cl_2$ order.

Why Fl_2 is least unexpected bond dissociation
⇒ Unexpected bond dissociation due to interelectronic repulsion because of its smaller size.

ii) Electron gain enthalpy.
> Process is exothermic
> Amount of energy released when an e^- is added in outermost shell of an isolated gaseous atom.

Why e^- gain enthalpy is high / more negative for Cl than Fl?

⇒ ~~In case of Fl, amount of energy~~

⇒ Due to smaller size of Fl, it has more interelectronic repulsion for incoming e^- whereas for Cl. due to its bigger size, interelectronic repulsion is size hence more negative.

> Hydration enthalpy. [exo]

Why Fl_2 is strongest ~~redox~~ oxidising agent

⇒ Due to its low bond dissociation enthalpy and high hydration enthalpy.

Electrode Potential

Tendency of an electrode to gain or loose e^-.

Q1. Given that the standard electrode potential (E°) of metals are

$K^+/K = -2.93\ V$, $Ag^+/Ag = 0.80\ V$, $Cu^{2+}/Cu = 0.34\ V$

$Mg^{2+}/Mg = -2.37\ V$.

$K > Mg > Cu > Ag$.

↓

Strongest reducing agent

Arrange these metals in an increasing order of their reducing power

Standard Hydrogen Electrode

Reference electrode to find electrode potential of unknown ~~half~~ electrode.

09/03/2024

What is the effect of dilution of specific conductance?

Ans k decreases with dilution as no. of ions per unit volume decrease.

8. The resistance of a conductivity cell containing 0.001M KCl solution at 298 K is 1500Ω. What is the cell constant if conductivity of 0.001 M KCl solution at 298 K is 0.146×10^{-3} S cm^{-1}.

Ans $R = 1500\,\Omega$

$k = 0.146 \times 10^{-3}$ S cm^{-1}

$$R = \rho \frac{l}{a}$$

$$\Rightarrow \frac{1}{\rho} \times R = \frac{l}{a}$$

$$\Rightarrow k \times R = \frac{l}{a}$$

$$\Rightarrow 0.146 \times 10^{-3} \times 1500 = \frac{l}{a}$$

$$\Rightarrow \frac{l}{a} = 2190\ cm^{-1}$$

146
×15
730
146
2190

Kohlrausch law of independent migrations of ions.

The law states that the limiting

Molar conductance of an electrolyte is the sum of molar conductance of its ions.

Application :

i) Molar conductance at given concentration λ_m^c

ii) Molar conductance at infinite dilution λ_m^∞

⇒ Means maximum molar conductance achieved after diluting a solution.

Molar conductance always increases with dilution.

For strong electrolyte, on dilution interionic attraction decreases hence ionic mobility increases

Cl^- K^+ Cl^-

Cl^- Cl^-

a) Calculation of Molar conductance at infinite solution for weak electrolytes.

$$\lambda_{H^+ OH^-}^{H_2O} = \lambda_{H^+ Cl^-}^{HCl} + \lambda_{Na^+ + OH^-}^{NaOH} - \lambda_{Na^+ Cl^-}^{NaCl}$$

λm° for NaCl, HCl and NaAC are 126.4, 425.9 and 91.0 $S\ cm^2\ mol^{-1}$ respectively. Calculate λ_0 for HA_c

Ans 425.9

91

516.9

- 126.4

390.5

2) Calculation of degree of dissociation of weak electrolytes.

α = degree of dissociation
Λ^c = molar conductance of solution at any conc.
Λ° = molar conductance of infinite dilution

$$\alpha = \frac{\Lambda^c_m}{\Lambda^\infty_m}$$

3) Calculation of dissociation constant of weak electrolytes

k = dissociation constant
α = degree of dissociation
C = concentration.

$$k = \frac{C\alpha^2}{1-\alpha}$$

$$\alpha = \frac{\Lambda^c_m}{\Lambda^\circ_m}$$

If $1 >> \alpha$

$$k = C\alpha^2$$

Q. 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^{-1}.

Ans. molarity

$C = 0.025 \text{ mol } L^{-1}$

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

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

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

$\Lambda^\circ_m (HCOOH) = 349.6 + 54.6$

$= 404.2 \text{ S cm}^2 \text{ mol}^{-1}$.

$$\alpha = \frac{\Lambda_m}{\Lambda^\circ_m}$$

$$= \frac{46.1}{404.2}$$

$$= 0.114$$

$$K = \frac{C\alpha^2}{1-\alpha}$$

$$= \frac{(0.025)(0.114)^2}{1 - 0.114}$$

$$= \frac{0.025 \times 12996}{0.886}$$

$$= \frac{0.00314900}{0.886}$$

$$= 3.55 \times 10^{-4} \text{ mol/L}$$

404.2 | 46.10 | 0.114

4042

5680

4042

16380

16188

212

1.000

0.114

0.886

0.114

1141

456

114

+114

12996

×25

64980

+25992

314900

10/05/2024.

8. Write the nerst eqn and calculate the emf of the given cell at 298 K.

Electrochemical cell

Cell which convert electrical energy to chemical energy.

~~Iron on electrode at anode~~.

8. Calculate the standard cell potential of galvanic cell in which the following reactions take place:

i) $2Cr(s) + 3Cd^{2+}(aq) \longrightarrow 2Cr^{3+}(aq) + 3Cd$.

Given. $E^\circ_{Cr^{3+}/Cr} = -0.74\,V$, $E^\circ_{Cd^{2+}/Cd} = -0.40\,V$

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

$E^\circ_{cell} = E^\circ_C - E^\circ_A$

$Cd^{2+} + 2e^- \longrightarrow Cd \quad -0.40\,V$.

$Cr^{3+} + 3e^- \longrightarrow Cr \quad -0.74\,V$.

$\downarrow$

$+0.74$.

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$= (-0.40) - (-0.74)$

$= +0.34\,V$

11) $Fe^{2+}(aq) + Ag^{+}(aq) \longrightarrow Fe^{3+}(aq) + Ag(s)$

$E^{\circ}_{cell} = E^{\circ}_{C} - E^{\circ}_{A}$

$= 0.80 - (-0.77)$

$= 0.03\ V.$

Nernst eqⁿ

$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n} \log \frac{[P]}{[R]}$$

Q. $Sn(s) / Sn^{2+}(0.050\ M) \,||\, H^{+}(0.020\ M) \,|\, H_2(g)\,(1\ bar) \,|\, Pt(s)$

Oxidation — Right - Reduction

Salt bridge

$Sn + 2H^{+} \longrightarrow Sn^{2+} + H_2$

$E^{\circ}_{cell} = E^{\circ}_{C} - E^{\circ}_{A}$

$= 0 - (-0.14)$

$= +0.14.$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.059}{2} \log \frac{[Sn^{2+}]}{[H^{+}]^2}$$

$$= 0.14 - 0.0295 \log \frac{0.05}{(0.02)^2}$$

Note → Stoichiometric coefficient are the power of conc in nerst eqⁿ.

$E_{cell} = 0.14 - 0.0295 \log \frac{0.05}{(0.02)^2}$

$= 0.14 - 0.0295 \log \frac{0.05}{0.02 \times 0.02}$

$= 0.14 - 0.0295 \log \frac{5}{4} \times \frac{10000}{100}$

$= 0.14 - 0.0295 (\log 5 + \log 10^2 - \log 4)$

$= 0.14 - 0.0295 (2.699 - 0.601)$

$= 0.14 - 0.0295 \times 2.097$

$= 0.14 - 0.06195$

$\therefore 0.09.$

$\log 4 = 0.602$

$\log 5 = 0.699$

* $\Delta G = -nFE^\circ$
* $E^\circ_{cell} = \frac{0.059}{n} \log K_c$

11/06/2024

$2Cr(s) + 3Cd^{2+}(aq) \longrightarrow 2Cr^{3+}(aq) + 3Cd(s)$

$\Delta G^\circ = -nFE^\circ_{cell}$

$= -6 \times 96500 \times 0.34$.

$= 196860$

$= 196.86$ kJ/mol.

$E^\circ_{cell} = E^\circ_C - E^\circ_A$

~~$= 0.74 - 1.0$~~

$= -0.4 - (-0.74)$

$= +0.34$

96500×34

3860000

2895

675500×6

$53\,0000$

$\Delta G = -nFE^\circ_{cell}$ — ①

Feasibility — $E^\circ_{cell} = +ve$

Nernst equation

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{n} \log \frac{[P]}{[R]} \quad \text{— ②}$$

At eq^m

$E_{cell} = 0$ $\quad R \longrightarrow P$

$K_{eq}/K_c \rightarrow \frac{[P]}{[R]}$

$$E^\circ_{cell} = \frac{0.059}{n} \log K_c \quad \text{— ③}$$

K_c is eq^m constant.

$$E_{cell} = E^\circ_{cell} - \frac{2.303RT}{nF} \log \frac{[P]}{[R]}.$$

R is gas constant [8.314 $JK^{-1}mol^{-1}$

T = 298 K

$$\frac{2.303RT}{F} = 0.059.$$

$$E_{cell} = E^\circ_{cell} - \frac{2.303\,RT}{nF} \log \frac{[P]}{[R]}$$

At eq^m $\quad E_{cell} = 0 \quad \frac{[P]}{[R]} = K_c$

$$E^\circ_{cell} = \frac{2.303RT}{nF} \log K_c \quad \text{--- (4)}$$

$$\Delta G = -nFE^\circ_{cell} \quad \text{--- (5)}$$

Put E°_{cell} value in eq^n (5)

$$\Delta G = -nF \times \frac{2.303RT}{nF} \log K_c$$

$$\Delta G = -2.303RT \log K_c$$

$$= -2.303 \frac{RT}{1} \log K_c$$

2.303
× 8.314
9212
2303
9212
18424
19.1 47.142

$$-19686.1 = -2.303 \times 8.314 \times 298 \log K_c$$

$$\frac{-19686.1}{2.303 \times 8.314 \times 298} = \log K_c$$

$$\Rightarrow 34.5 = \log K_c$$

Taking antilog on both side.

$$K_c = \text{antilog } 34.5$$

ii) $Fe^{2+}(aq) + Ag^{+}(aq) \longrightarrow Fe^{3+}(aq) + Ag(s)$

$E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$

$= 0.80 - (0.77)$

$= 0.03\ V.$

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

$= -1 \times 96500 \times 0.03$

$= 2895\ J/mol.$

~~$\log K_c = 1 \times 0.03$~~

$\Delta E^{\circ}_{cell} = -\frac{0.059}{n} \log K_c$

$\log K_c = \frac{n E^{\circ}_{cell}}{0.059}$

$= \frac{1 \times 0.03}{0.059}$

$= 0.508$

Taking antilog on both side

$K_c = 3.22$

Q. Represent the cell in which the following reaction takes place. The value of E° for the cell is 1.260 V. What is the value of E_{cell}?

$3(d\ (s) + 2Al^{3+}(0.01M) \rightarrow 2Al(s) + 3Cd^{2+}(0.1M).$

$Al(s) / Cd^{2+}(0.1M) // Al^{3+}(0.01M) / Cd(s)$

$2Al(s) + 3Cd^{2+}(0.1M) \rightarrow 3Cd(s) + 2Al^{3+}(0.01M)$

$E^\circ_{cell} = E^\circ_{cell} - \frac{0.059}{n} \log \frac{[Al^3]^2}{[Cd^2]^3}$ ~~$E^\circ_{cell} = E^\circ_c - E^\circ_a$~~

$E_{cell} = 1.26 - \frac{0.059}{6} \log \frac{(0.01)^2}{(0.1)^3}$

$= 1.26 - \frac{0.059}{6} \frac{(0.01)^2}{(0.1)^3} \frac{(10^{-2})^2}{(10^{-1})^3}$

$= 1.26 - 0.0295 \log \frac{10^{-3}}{} \times 10^7$

$= 1.26 - 0.0295 \log 10$

$= 1.26 - 0.0295 (-1)$

$= 1.25$

Faraday's 1st Law.

It states that the amount of a substance that is liberated or deposited at electrode is $\propto$ proportional to the quantity of electricity passed through the electrodeyte

$w \propto Q \qquad m \propto q.$

$m = \frac{MIt}{nx96500}$

$m = ZIt$

$m = \frac{M}{n} \times \frac{I x t}{96500}$

where m - mass
M → atomic mass
n → valency
I → current
t → time.

13/05/24.

2nd Faraday Law.

When the same quantity of electricity is passed through electrolytic solution, the weigh of different substances produced at the electrodes ratios of their equivalent masses.

$$\frac{M_{Ag}}{M_{Cu}} = \frac{\frac{M}{n} \times \frac{I \times t}{96500}}{\frac{M}{n} \times \frac{I \times t}{96500}}$$

$$\frac{M_{Ag}}{M_{Cu}} = \frac{\text{eq. mass. Ag}}{\text{eq. mass. Cu}}$$

Batteries

Primary batteries → reaction occurs only once and after use over a period of time battery becomes dead & cannot be reused again.

Leclanche cell is used commonly in our transistors & clocks.

The cell consists of a zinc container that also acts as anode & the cathode is a carbon (graphite) rod surrounded by powdered manganese dioxide & carbon.

The space b/w the electrodes is filled by a moist paste of ammonium chloride (NH_4Cl) & zinc chloride ($ZnCl_2$).

Ammonia produced in the reaction forms a complex with Zn^{2+} to give $[Zn(NH_3)_4]_{2+}$. The cell has a potential of nearly 1.5 V.

13/09/2024

Anode = $Zn(s) \longrightarrow Zn^{2+} + 2e^-$

Cathode = $MnO_2 + NH_4^- + e^- \longrightarrow MnO(OH) + NH_3$

Mercury cell suitable for low current devices like hearing aids, watches, etc.

- Consists of zinc - mercury amalgam as anode & a paste of HgO & carbon as the cathode.

The electrolyte is a paste of KOH & ZnO. The overall reaction is represented by

$Zn(Hg) + HgO(s) \rightarrow ZnO(s) + Hg(l)$.

The cell potentials is approx. 1.35 V and remains constant during its life as the overall reaction does not involve any ion in solution whose conc. can change during its lifetime.

Role of $ZnCl_2$ in Leclanche cell.

Ammonia produced in the reaction forms a complex with Zn^{2+} to give $[Zn(NH_3)_4]^{2+}$. The cell has a potential of nearly 1.5 V.

Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.

$$H^+ + e^- \longrightarrow \frac{1}{2} H_2$$

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{1} \log \frac{1}{[H^+]}$$

$$= 0 - 0.059 \log \frac{1}{10^{-10}}$$

$$= -0.059 \log 10^{20}$$

$$= -0.059 \times 10 \log 10$$

$$E_{cell} = 0.59$$

$pH = 10 = -\log [H^+]$

$-10 = \log [H^+]$

Taking antilog on both side

antilog $(-10) = [H^+]$

$10^{-10} = [H^+]$

$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$ in aqueous solution

get hydrated

being smaller most highly hydrated become heavy hence ionic mobility decrease / conductivity ↓↓

Cs^+ biggest cation & hence least hydrated hence ionic mobility increases.

17. Soln of 2 electrolytes A & B are diluted. The Λ_m of B increases 1.3 times while that of A increases 25 times. Which of the two is a strong electrolyte? Justify your ans.

Ans B.