

CHEMISTRY

50=Practical

Internal assessment=20

Final exam=30



Theory = 100

Internal assessment=20

Final exam=80

Four teachers 1h to each

- PHYSICAL CHEMISTRY ; 1h, =16
- ENVIRONMENTAL CHEMISTRY; 1h=16
- INORGANIC CHEMISTRY; 1h=16
- ORGANIC CHEMISTRY; 1h=16

Books

- 1. A Text book of Engineering Chemistry; By Amala Shrestha and group**
- 2. A Text Book of Engineering Chemistry; By R. K. Sharm & group**

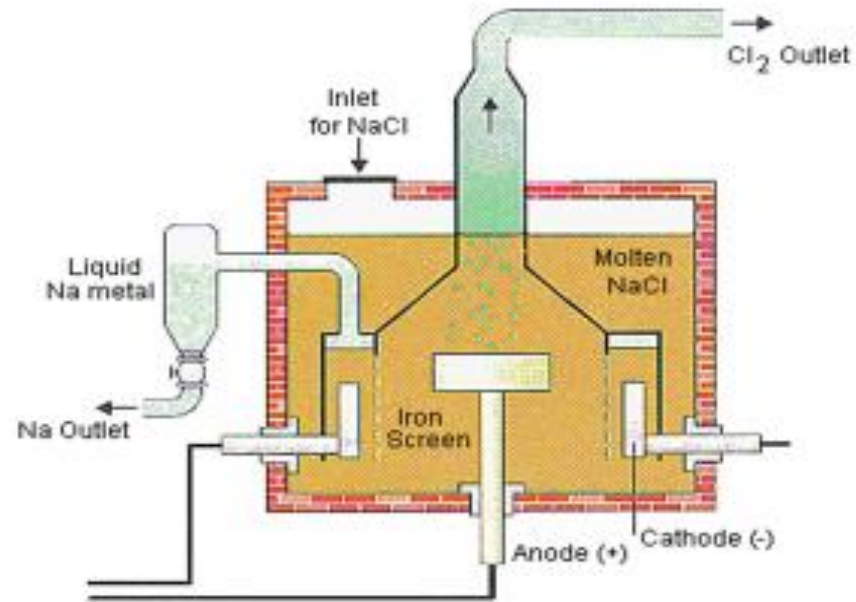
For Physical Chemistry

- 1. A Selected Topics in Physical Chemistry; By M. K. Sthapit**
- 2. Essential of Physical Chemistry; By B. S. Bahl & G.D Tuli**
- 2. Principles of Physical Chemistry; By Maron And Prutton**

Physical Chemistry

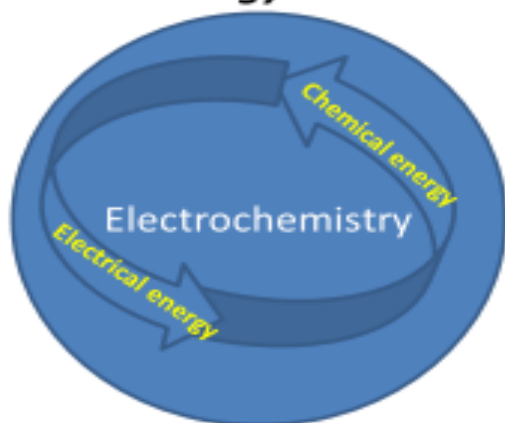
- Electrochemistry
- Buffer solution
- catalysis

Electrochemistry



Electrochemistry

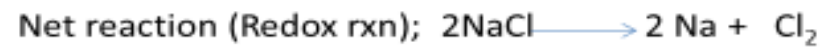
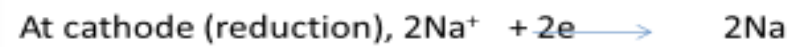
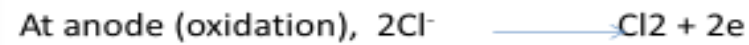
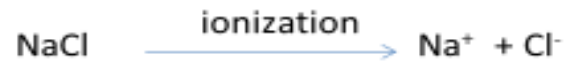
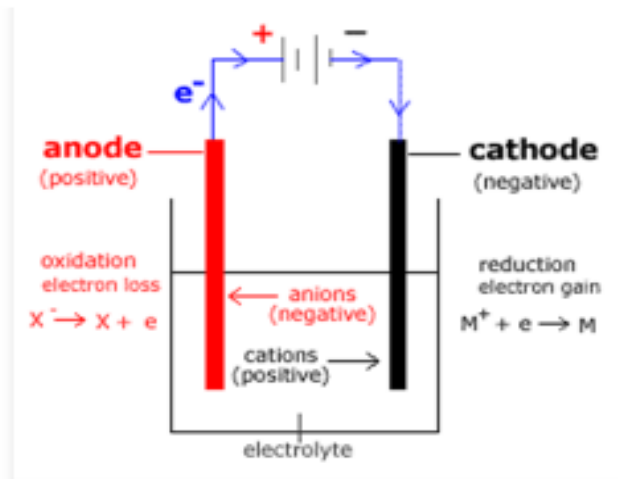
Conversion of electrical energy in to chemical energy and vice versa



1. Chemical reaction due to electric energy; *Electrolytic cell*

2. Electric energy due to chemical reaction; *Galvanic cell*

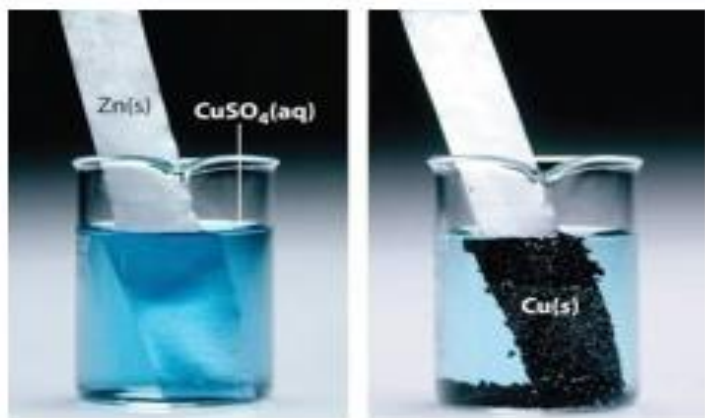
1. Electrolytic cell



Electrolyte & non electrolyte
Strong & weak electrolyte

1. Electroplating
2. Electrometallurgy
3. Nanofiber fabrication(Electrospinning)

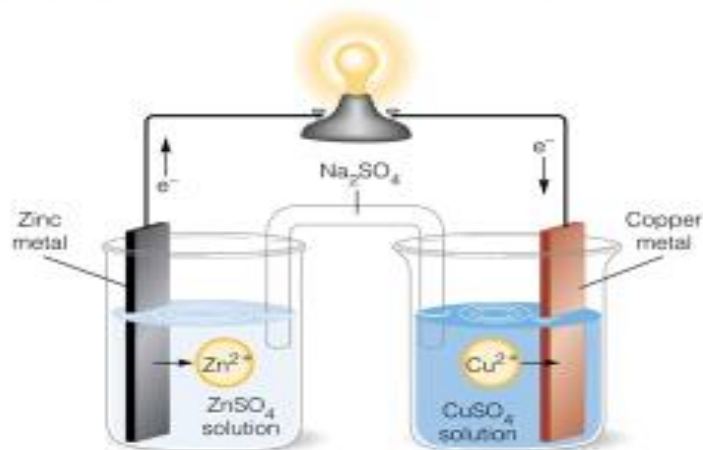
Concept of galvanic cell



Ionic reaction = redox rxn



Chemical Energy into heat energy

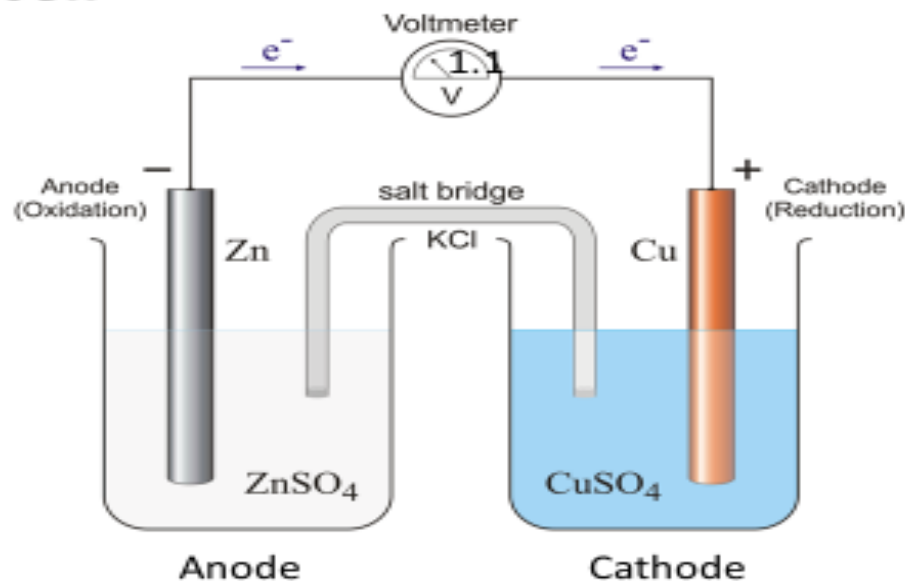


Chemical Energy into electrical energy

2. Galvanic cell or voltaic cell

A. Zn-Cu cell

What is Daniell cell?



At anode (oxidation)

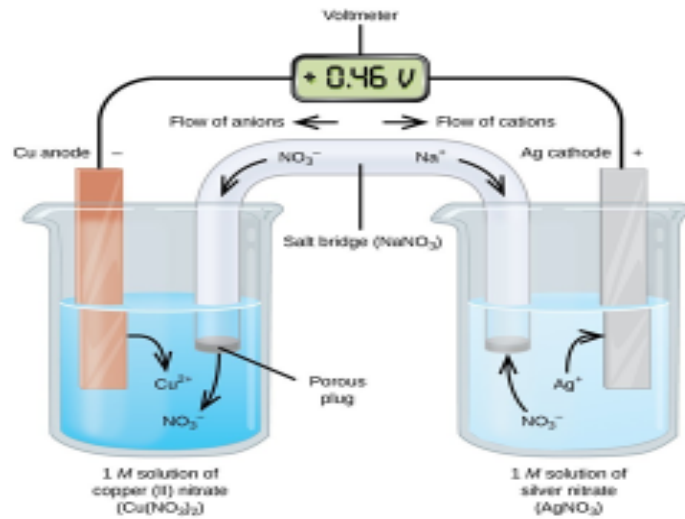


At cathode (reduction)



Net cell reaction; $\text{Zn} + \text{Cu}^{++} \longrightarrow \text{Zn}^{++} + \text{Cu}$, $E^{\circ}_{\text{cell}} = 1.1\text{V}$

B. Cu-Ag cell



What is Standard Condition of cell?

At anode (oxidation)



At cathode (reduction)



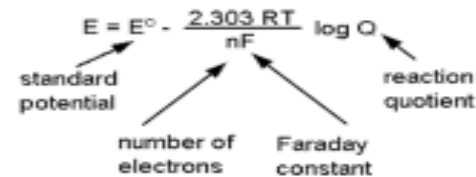
Standard condition of cell

1. Room temperature is 25 C
2. Concentration of used electrolyte is 1M
3. If gas is used as electrode, its pressure should be 1 atm

EMF measured at this condition is, E^0_{cell}

Change of any condition; , E_{cell}

Nernst equation



Electrochemical Series

Elements	Electrode Reaction	E°_{red} (volts)
<i>Oxidised Form + ne⁻ → Reduced Form</i>		
Li	$Li^{+}(aq) + e^{-} \longrightarrow Li(s)$	-3.05
K	$K^{+}(aq) + e^{-} \longrightarrow K(s)$	-2.93
Ba	$Ba^{2+}(aq) + 2e^{-} \longrightarrow Ba(s)$	-2.90
Ca	$Ca^{2+}(aq) + 2e^{-} \longrightarrow Ca(s)$	-2.87
Na	$Na^{+}(aq) + e^{-} \longrightarrow Na(s)$	-2.71
Mg	$Mg^{2+}(aq) + 2e^{-} \longrightarrow Mg(s)$	-2.37
Al	$Al^{3+}(aq) + 3e^{-} \longrightarrow Al(s)$	-1.66
Zn	$Zn^{2+}(aq) + 2e^{-} \longrightarrow Zn(s)$	-0.76
Cr	$Cr^{3+}(aq) + 3e^{-} \longrightarrow Cr(s)$	-0.74
Fe	$Fe^{2+}(aq) + 2e^{-} \longrightarrow Fe(s)$	-0.44
	$H_2O(l) + e^{-} \longrightarrow \frac{1}{2}H_2(g) + OH^{-}(aq)$	-0.41
Cd	$Cd^{2+}(aq) + 2e^{-} \longrightarrow Cd(s)$	-0.40
Pb	$PbSO_4(s) + 2e^{-} \longrightarrow Pb(s) + SO_4^{2-}(aq)$	-0.31
Co	$Co^{2+}(aq) + 2e^{-} \longrightarrow Co(s)$	-0.28
Ni	$Ni^{2+}(aq) + 2e^{-} \longrightarrow Ni(s)$	-0.25
Sn	$Sn^{2+}(aq) + 2e^{-} \longrightarrow Sn(s)$	-0.14
Pb	$Pb^{2+}(aq) + 2e^{-} \longrightarrow Pb(s)$	-0.13
H₂	$2H^{+} + 2e^{-} \longrightarrow H_2(g)$ (standard electrode)	0.00
Cu	$Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$	+0.34
I ₂	$I_2(s) + 2e^{-} \longrightarrow 2I^{-}(aq)$	+0.54
Fe	$Fe^{3+}(aq) + e^{-} \longrightarrow Fe^{2+}(aq)$	+0.77
Hg	$Hg_2^{2+}(aq) + 2e^{-} \longrightarrow 2Hg(l)$	+0.79
Ag	$Ag^{+}(aq) + e^{-} \longrightarrow Ag(s)$	+0.80
Hg	$Hg^{2+}(aq) + 2e^{-} \longrightarrow Hg(l)$	+0.85
N ₂	$NO_3^{-} + 4H^{+} + 3e^{-} \longrightarrow NO(g) + 2H_2O$	+0.97
Br ₂	$Br_2(l) + 2e^{-} \longrightarrow 2Br^{-}(aq)$	+1.08
O ₂	$O_2(g) + 2H_2O^{+}(aq) + 2e^{-} \longrightarrow 3H_2O$	+1.23
Cr	$Cr_2O_7^{2-} + 14H^{+} + 6e^{-} \longrightarrow 2Cr^{3+} + 7H_2O$	+1.33
Cl ₂	$Cl_2(g) + 2e^{-} \longrightarrow 2Cl^{-}(aq)$	+1.36
Au	$Au^{3+}(aq) + 3e^{-} \longrightarrow Au(s)$	+1.42
Mn	$MnO_4^{-}(aq) + 8H_2O^{+}(aq) + 5e^{-} \longrightarrow Mn^{2+}(aq) + 12H_2O(l)$	+1.51
F ₂	$F_2(g) + 2e^{-} \longrightarrow 2F^{-}(aq)$	+2.87

(a) Tendency for oxidation to occur
(b) Power as reducing agent

(a) Tendency for reduction to occur
(b) Power as oxidising agent

Project for you

- Application of electrochemistry in Comp Eng

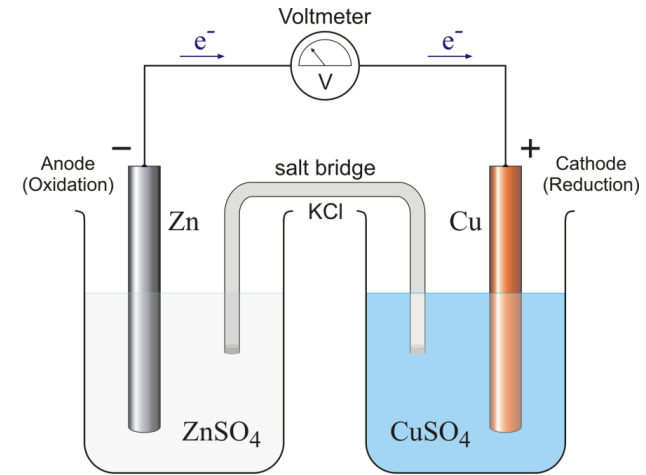
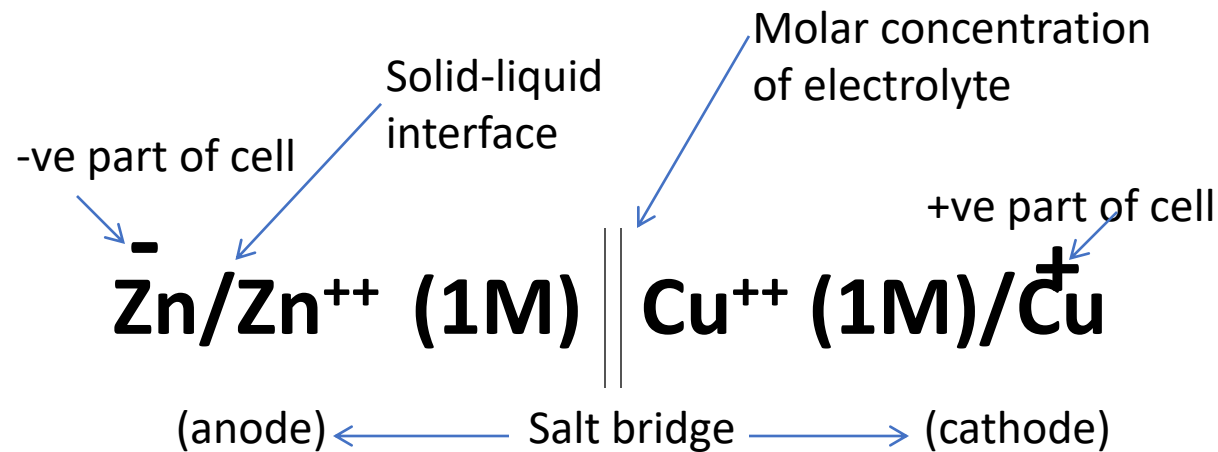
Hints; memory card
energy storage device

Question for examination

- 1 . Difference between electrolytic cell and galvanic cell
2. What is Daniell cell?
3. What is standard cell?

Cell notation

Cell notation in electrochemistry is a shorthand way of expressing a complete galvanic cell diagram in a single line

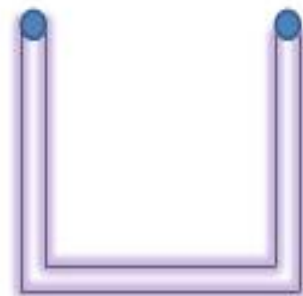
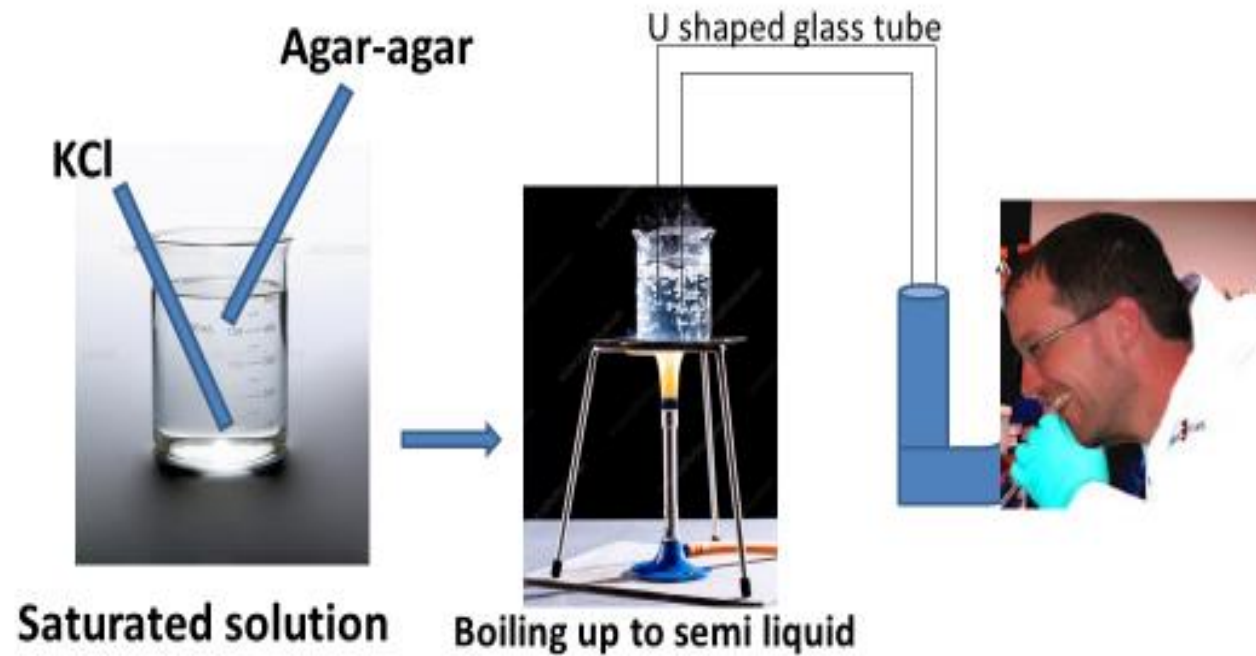


Daniell cell

Salt bridge

- U-shaped glass tube filled with gel of agar-agar with large amount of inert salt (KCl , NaNO_3 , NH_4NO_3)
- Minimize liquid-liquid junction potential (**Potential arises between two solutions when they are in contact with each other**)
- Complete electric circuit
- Maintain flow of ions by providing aqueous medium and maintain charge balance in each half cell

Preparation of Salt bridge



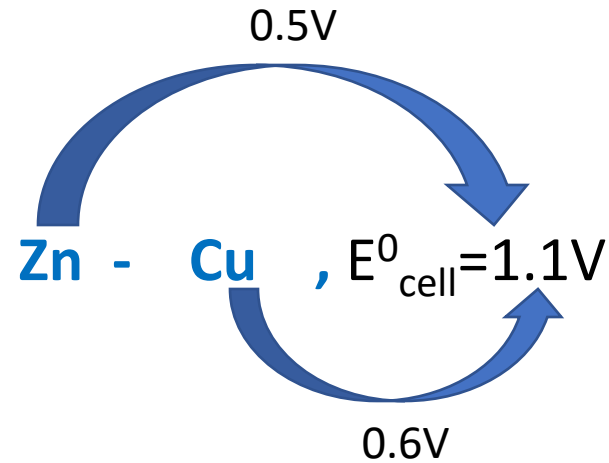
U-shaped glass tube filled with gel of Agar-
Agar with large amount of inert salt

Single electrode potential

A. Zn-Cu cell, $E^0_{\text{cell}} = 1.1\text{V}$

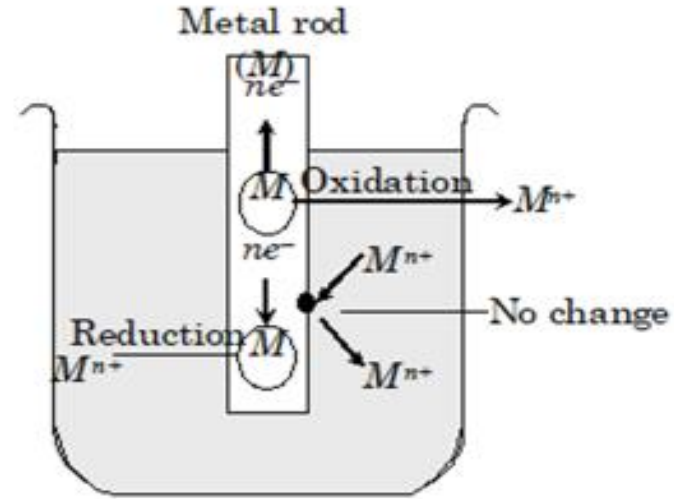
B. Cu-Ag cell, $E^0_{\text{cell}} = 0.46\text{V}$

?

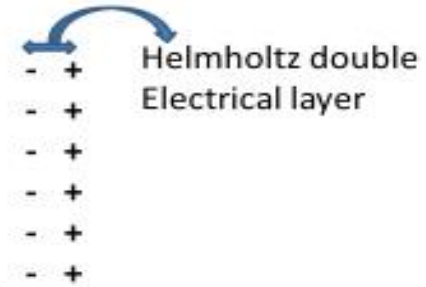
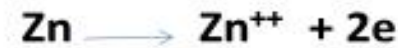
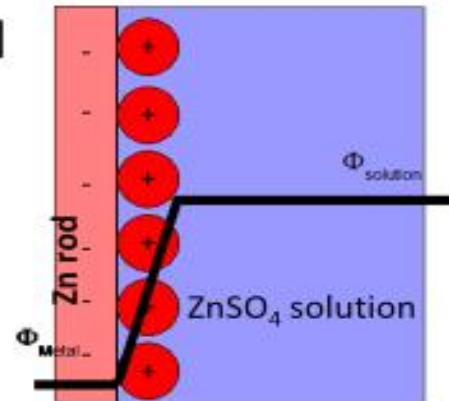


Now single electrode potential of Zn electrode is 0.5V
And that of Cu is 0.6V

Origin of single electrode potential



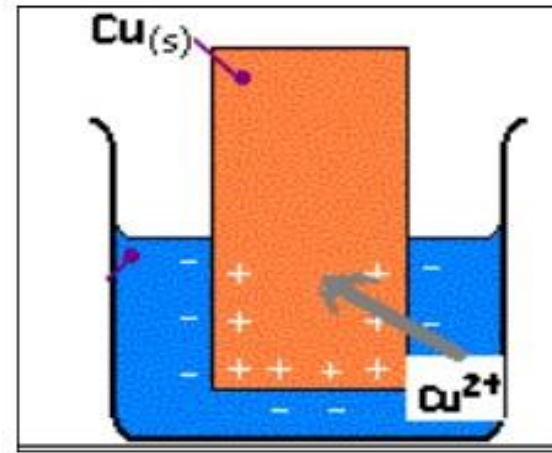
Case I



Standard single electrode potential

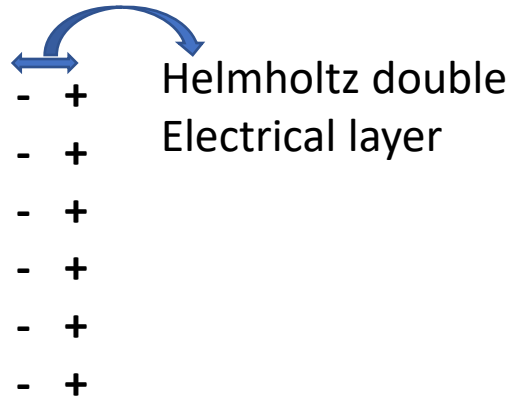
Cu rod in to M CuSO₄ solution at 25 degree centigrade

When Cu is placed in the solution of CuSO₄, the copper ions in the solution gets deposited over the metal and hence the metal becomes positively charged. The positively charged metal attracts the negatively charged sulphate ions in the solution and hence a doubly charged layer (Helmholtz electrical double layer) is formed near the metal.



When any metallic rod is partially dipped into its soluble salt solution, there is formation of oppositely charged double electrical layer called Helmholtz double layer. Due to this oppositely charged layer, certain potential difference should be developed across the solid-liquid interface which is called single electrode potential. If it is measured at standard condition, it is called standard single electrode potential

Measurement of single electrode potential



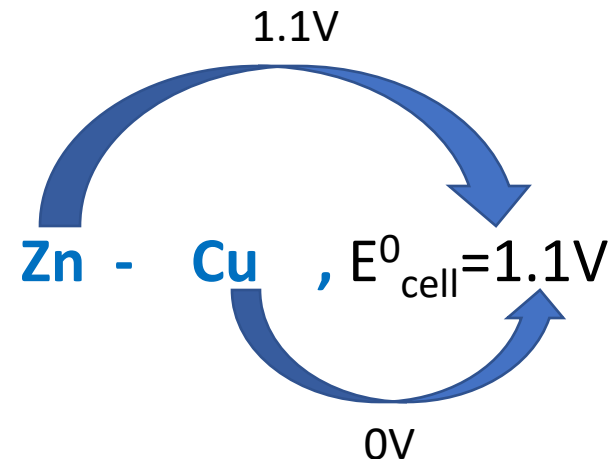
Size of electron and atom in $\text{\AA}^0 = 10^{-10} \text{ m}$

What is the size of material experimentally fabricated and measured ?

10^{-9} m = nanometer = nano technology

There is no any instrument to measure this potential difference between these opposite charged layer up to now

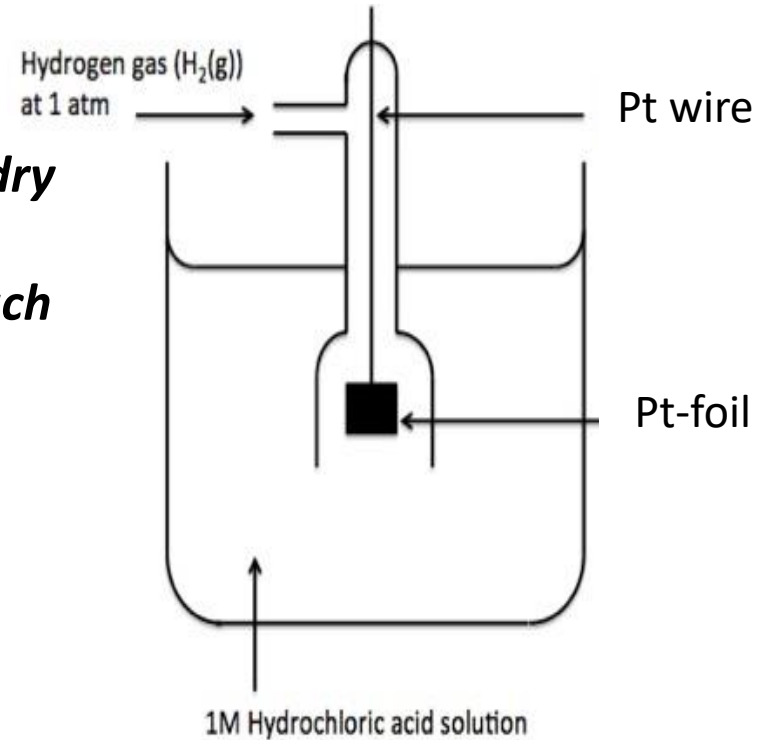
Therefore need some reference
Electrode having 0V electrode potential



Normal hydrogen electrode (NHE)

Normal hydrogen electrode consists of 1M HCl solution in to a beaker through which pure and dry H₂ gas is passed at 1 atm. To introduce electric current through electrode system, inert metal such as Pt-foil is dipped into HCl solution. This simple system as shown in this figure at 25 degree centigrade is called NHE or SHE.

Single electrode potential of NHE is near about 0V. Therefore, it can be used as reference electrode to Measure the single electrode potential of any electrode

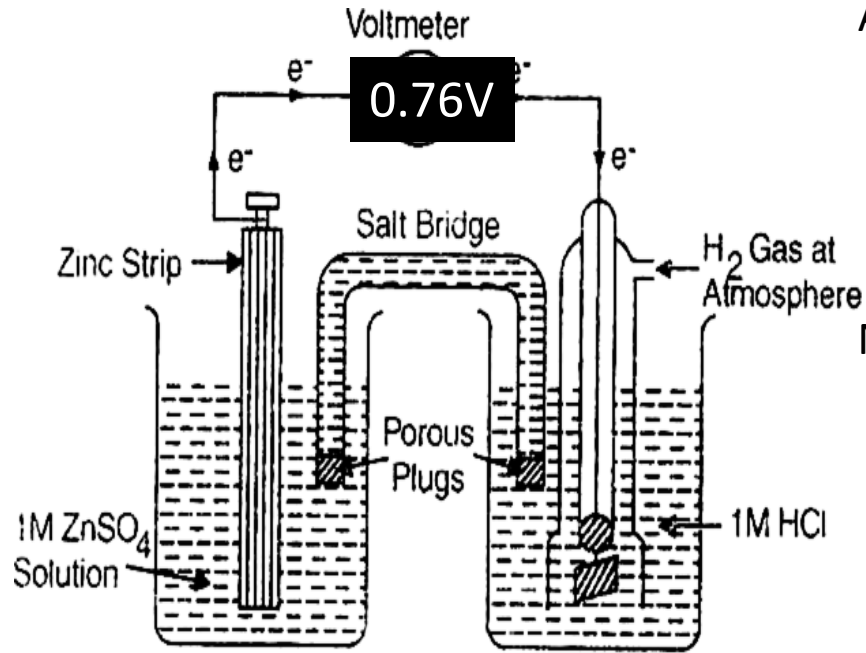


Question for examination

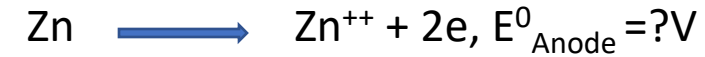
- 1 . What is cell notation?
2. Write down the function of salt bridge.
3. What is standard single electrode potential
4. What is NHE or SHE?

- Standard single electrode potential
- No instrument to measure SSEP
- Need reference electrode having 0 V SSEP
- NHE is found as RE having 0 V electrode pot.
- Should make complete galvanic cell with NHE
- NHE anode, cell notation is; $\text{Pt}, \text{H}_2(1\text{atm})/\text{H}^+(1\text{M})//$
- NHE anode, cell notation is; $//\text{H}^+(1\text{M})/\text{H}_2(1\text{atm})\text{Pt}^+$

Measurement of single electrode potential of Zn electrode



At anode (oxidation)



At cathode (reduction)



$$\text{Since, } E^0_{\text{A}} + E^0_{\text{C}} = 0.76\text{V}$$

$$\text{or, } E^0_{\text{A}} + 0\text{V} = 0.76\text{V}$$

It is standard Single oxidation pot and given by

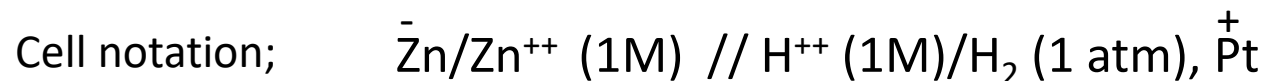
$$E^0_{\text{Zn/Zn}^{++}} = 0.76\text{v}$$

It is in terms of reduction

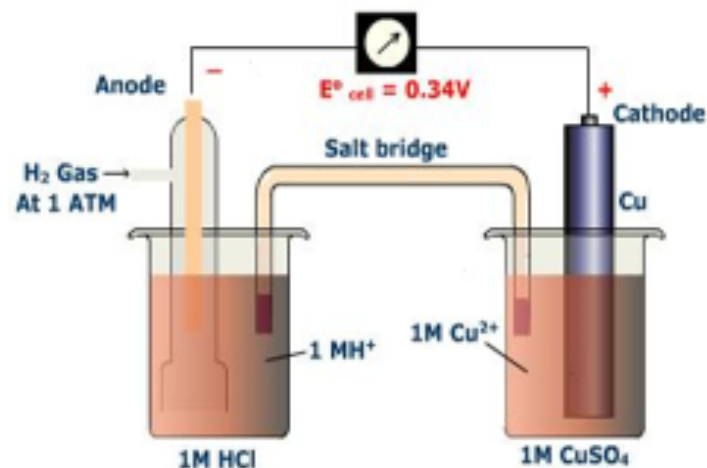
$$E^0_{\text{Zn}^{++}/\text{Zn}} = -0.76\text{v}$$

For any electrode;

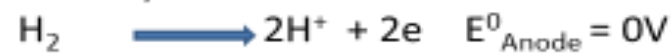
Standard single reduction electrode potential = - Standard single oxidation electrode potential



Measurement of single electrode potential of Cu electrode



At Anode (oxidation)



At Cathode (reduction)



$$\text{Since, } E^0_{\text{A}} + E^0_{\text{C}} = 0.34\text{V}$$

$$\text{or, } E^0_{\text{C}} = 0.34\text{V}$$

It is standard Single reduction pot and given by

$$E^0_{\text{Cu}^{++}/\text{Cu}} = 0.34\text{V}$$

It is in terms of oxidation

$$E^0_{\text{Cu}/\text{Cu}^{++}} = -0.34\text{V}$$

Cell notation; $\text{Pt, H}_2(1\text{atm})/\text{H}^+(1\text{M}) // \text{Cu}^{++}(1\text{M})/\text{Cu}$

Foe Ag electrode, it is cathode $E^0_{\text{Ag}^+/\text{Ag}} = 0.80\text{V}$ (reduction potential)

$$E^0_{\text{Ag}/\text{Ag}^+} = -0.80\text{V} \text{ (oxidation potential)}$$

Electrochemical series (ECS)

Measurement of St. Sing. Elect. Pot of each electrode
With the help of NHE

Change all pot in terms of reduction and arrange them
according to their increasing St Sing Red Elect Pot.

Gives series called ECS)

Zn electrode= Anode = oxidation

$$= E^0_{\text{Zn/Zn}^{++}} = 0.76\text{V}, E^0_{\text{Zn}^{++}/\text{Zn}} = -0.76\text{V}$$

Cu electrode= Cathode, $E^0_{\text{Cu}^{++}/\text{Cu}} = 0.34\text{V}$

Ag electrode= Cathode, $E^0_{\text{Ag}^+/\text{Ag}} = 0.80\text{V}$

Fe electrode= Anode = oxidation

$$= E^0_{\text{Fe/Fe}^{++}} = 0.44\text{V}, E^0_{\text{Cu}^{++}/\text{Cu}} = -0.44\text{V}$$

Li	=	- 3.05V
.		
.		
.		
Zn	=	- 0.76V
Fe	=	- 0.44V
H	=	0V
Cu	=	+0.34
Ag	=	+0.80V
.		
.		
.		
.F2	=	+2.87V

ECS

Application of ECS

1. Strength of oxidizing & reducing agent

Moving from top to down in ECS, oxidizing strength is increased where as reducing strength is decreased.

2. Construction of galvanic cell

Electrode having low value of standard reduction potential is always anode and that which has high value of standard reduction potential is always cathode.

Q. Write down the cell notation of given Combination

$$E_{X/X^+}^0 = -0.64V \quad E_{M^{++}/M}^0 = +0.24V,$$

Solution, $M/M^{++} // X^+/X$

Elements	Electrode Reaction	E_{red}^0 (volts)
<div style="display: flex; justify-content: space-between;"> (a) Tendency for oxidation to occur (b) Power as reducing agent Increase (a) Tendency for reduction to occur (b) Power as oxidising agent </div>		
<div style="text-align: center;">Oxidised Form + $ne^- \longrightarrow$ Reduced Form</div>		
Li	$Li^+(aq) + e^- \longrightarrow Li(s)$	-3.05
K	$K^+(aq) + e^- \longrightarrow K(s)$	-2.93
Ba	$Ba^{2+}(aq) + 2e^- \longrightarrow Ba(s)$	-2.90
Ca	$Ca^{2+}(aq) + 2e^- \longrightarrow Ca(s)$	-2.87
Na	$Na^+(aq) + e^- \longrightarrow Na(s)$	-2.71
Mg	$Mg^{2+}(aq) + 2e^- \longrightarrow Mg(s)$	-2.37
Al	$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$	-1.66
Zn	$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$	-0.76
Cr	$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$	-0.74
Fe	$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	-0.44
	$H_2O(l) + e^- \longrightarrow \frac{1}{2}H_2(g) + OH^-(aq)$	-0.41
Cd	$Cd^{2+}(aq) + 2e^- \longrightarrow Cd(s)$	-0.40
Pb	$PbSO_4(s) + 2e^- \longrightarrow Pb(s) + SO_4^{2-}(aq)$	-0.31
Co	$Co^{2+}(aq) + 2e^- \longrightarrow Co(s)$	-0.28
Ni	$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$	-0.25
Sn	$Sn^{2+}(aq) + 2e^- \longrightarrow Sn(s)$	-0.14
Pb	$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$	-0.13
H₂	$2H^+ + 2e^- \longrightarrow H_2(g)$ (standard electrode)	0.00
Cu	$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.34
I ₂	$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$	+0.54
Fe	$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	+0.77
Hg	$Hg_2^{2+}(aq) + 2e^- \longrightarrow 2Hg(l)$	+0.79
Ag	$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.80
Hg	$Hg^{2+}(aq) + 2e^- \longrightarrow Hg(l)$	+0.85
N ₂	$NO_3^- + 4H^+ + 3e^- \longrightarrow NO(g) + 2H_2O$	+0.97
Br ₂	$Br_2(aq) + 2e^- \longrightarrow 2Br^-(aq)$	+1.08
O ₂	$O_2(g) + 2H_3O^+(aq) + 2e^- \longrightarrow 3H_2O$	+1.23
Cr	$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$	+1.33
Cl ₂	$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$	+1.36
Au	$Au^{3+}(aq) + 3e^- \longrightarrow Au(s)$	+1.42
Mn	$MnO_4^-(aq) + 8H_3O^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 12H_2O(l)$	+1.51
F ₂	$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$	+2.87
<div style="display: flex; justify-content: space-between;"> Increase </div>		

3. Calculation of EMF of cell at standard condition

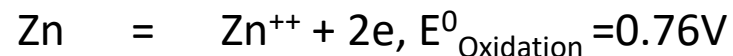
~~$$E^0_{\text{cell}} = E^0_{\text{C}} - E^0_{\text{A}}$$~~

$$E = E^0 - \frac{2.303 RT}{nF} \log Q$$

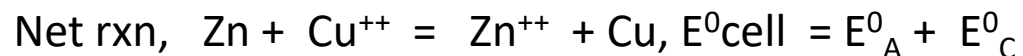
standard potential \rightarrow E^0
number of electrons \rightarrow n
Faraday constant \rightarrow F
reaction quotient \rightarrow Q

$Q = [\text{Zn}^{++}]/[\text{Cu}]$

Use; At anode (oxidation)



At cathode (reduction)



4. Spontaneity of redox reaction

If EMF of cell is +ve, the given redox rxn towards forward direction is spontaneous

And if it is -ve, the redox rxn is nonspontaneous (not feasible)

5. Slow and fast redox reaction

Larger the distance between two electrode faster will be the redox reaction



6. Calculation of ΔG and K_{eq} of reaction

$$\Delta G = -nFE$$

$$\Delta G = -RT \ln K_{\text{eq}}$$

7. Only those metal having upper position in ECS can displace another metal from soluble salt which is below in the ECS

Fe vessel can not be used to store CuSO_4 solution but it can be used to store ZnSO_4 solution

8. Metals above the Hydrogen in ECS can produce H_2 gas from nonoxidising acid but metals below H_2 could not

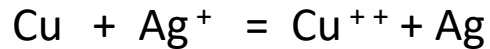
Calculation of emf of cell or electrode potential at nonstandard condition (Nernst equation)

Standard condition; (1) Temperature 25° C

(2) Concentration of electrolyte 1M

(3) If gas is used as electrode, its pressure 1 atm

For given cell, calculate the emf at standard condition

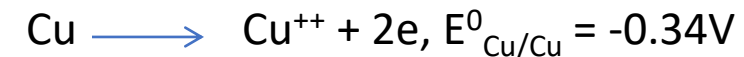


Given, $E^0_{\text{Cu}/\text{Cu}^{++}} = -0.34\text{V}$, $E^0_{\text{Cu}^{++}/\text{Cu}} = +0.34\text{V}$

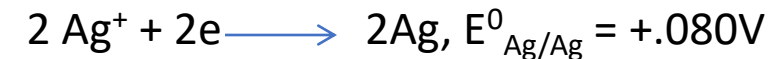
$E^0_{\text{Ag}^+/\text{Ag}} = +0.80\text{V}$

$$\begin{aligned}\text{Solution; } E^0_{\text{Cell}} &= E^0_{\text{C}} + E^0_{\text{A}} \\ E^0_{\text{Cell}} &= 0.80 - 0.34 \\ &= 0.46\text{V}\end{aligned}$$

At anode (oxidation)



At cathode (reduction)



$$E_{\text{cell}} = E^0_{\text{cell}} - \frac{RT}{nF} \ln Q \quad Q = [\text{Cu}^{++}] [\text{Ag}]^2 / [\text{Cu}] [\text{Ag}^+]^2$$