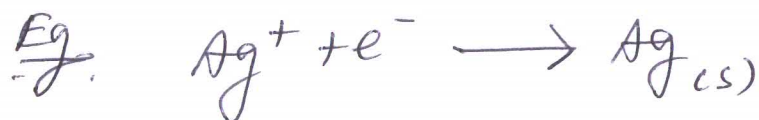


①

Electrodes: (a) Galvanic cell.

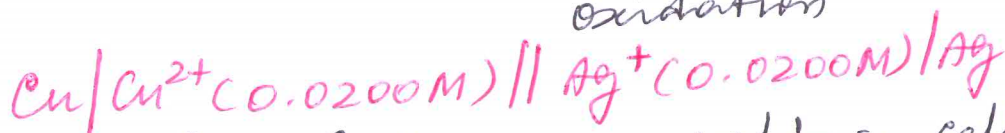
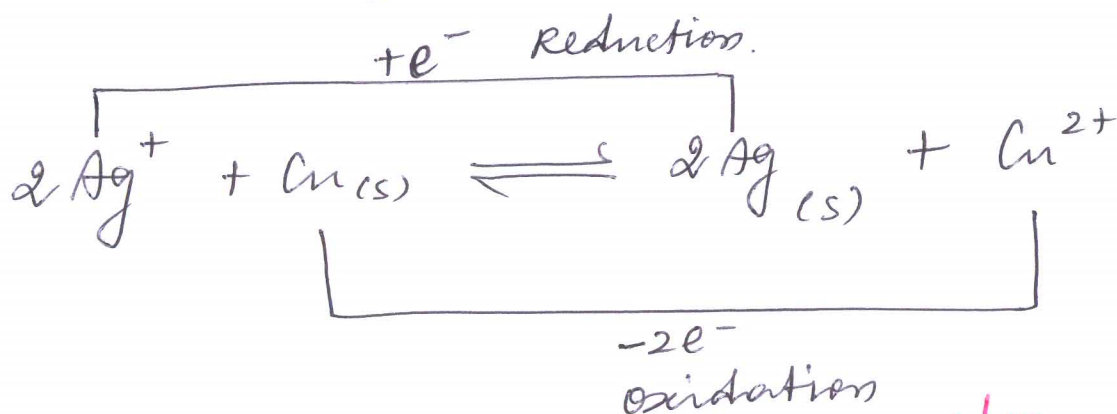
Cathode:

Reduction occurs here.

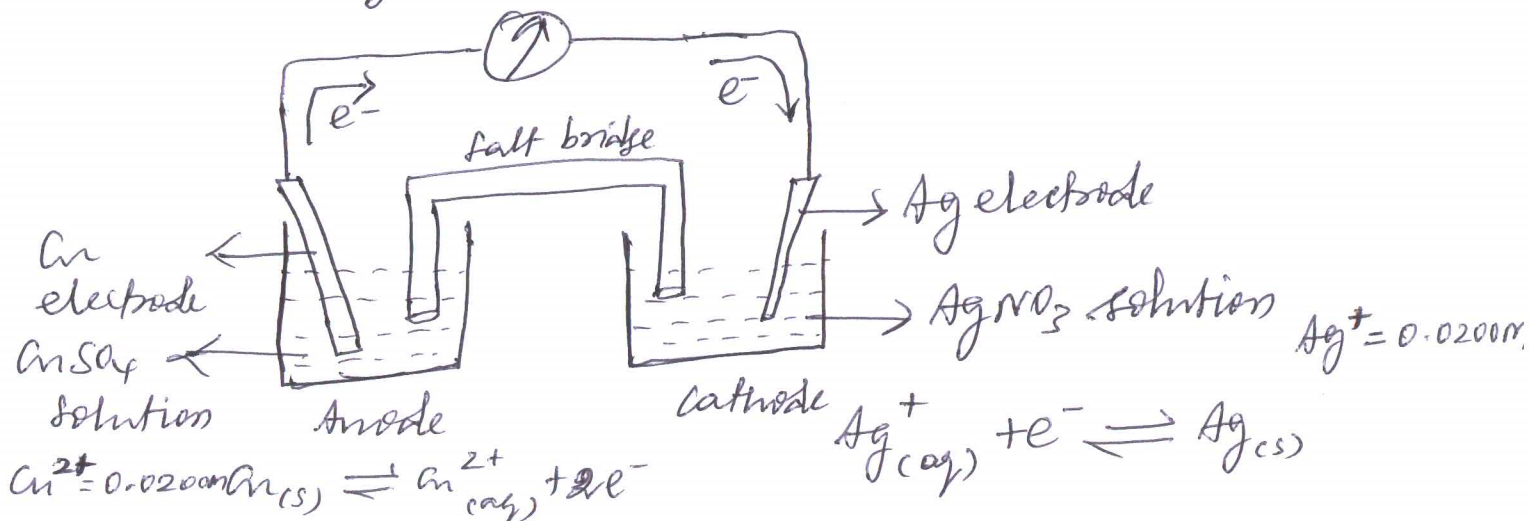


Anode:

oxidation occurs here.

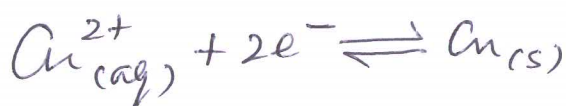
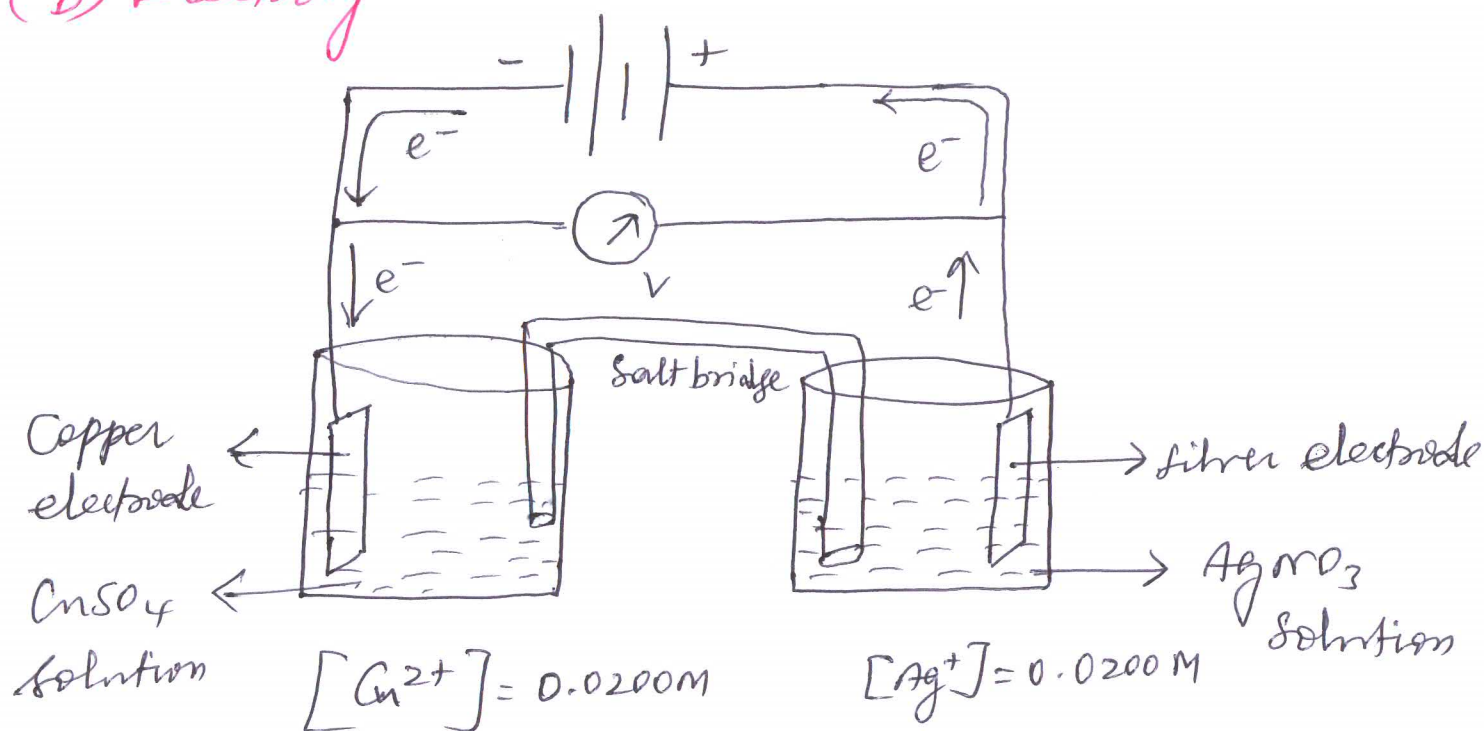


Eg. Galvanic or voltaic cells

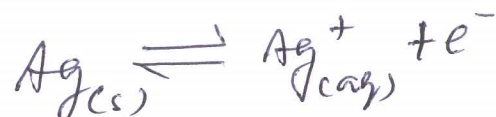


(2)

(b) Electrolytic cells:



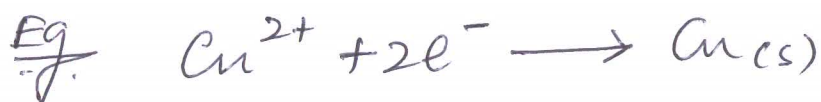
Cathode



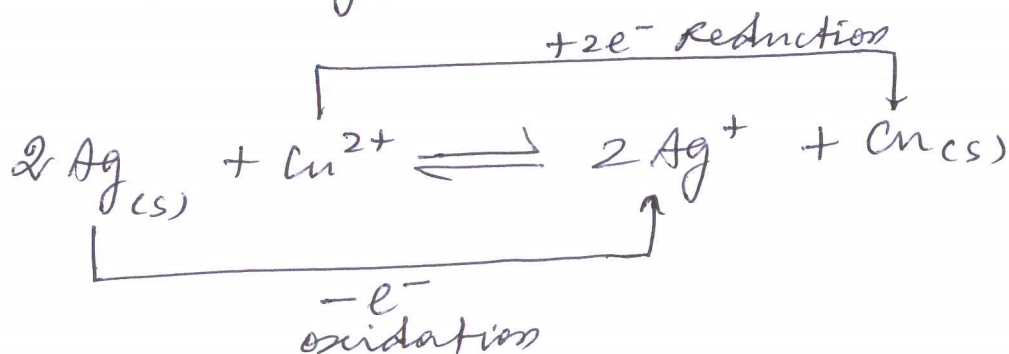
Anode.



Cathode: Reduction occurs here.



Anode: Oxidation occurs here.



Chemical Reversible Reactions :

③

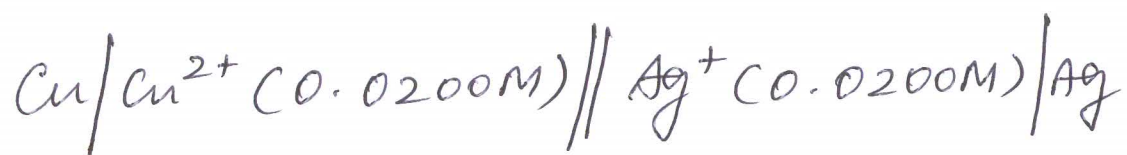
Chemical reversible reactions are those which when forced to proceed in the opposite directions (add electrical energy by reversing current) results in the original reactants. (no new reactions appear Eg. rechargeable batteries).

Chemical Irreversible Reaction :

Chemical irreversible reaction are those which when forced to proceed in the opposite (nonspontaneous) direction, do not result in the original reactants but form new products (Eg. non rechargeable batteries).

Schematic Representation of cells:-

(A)



By convention,

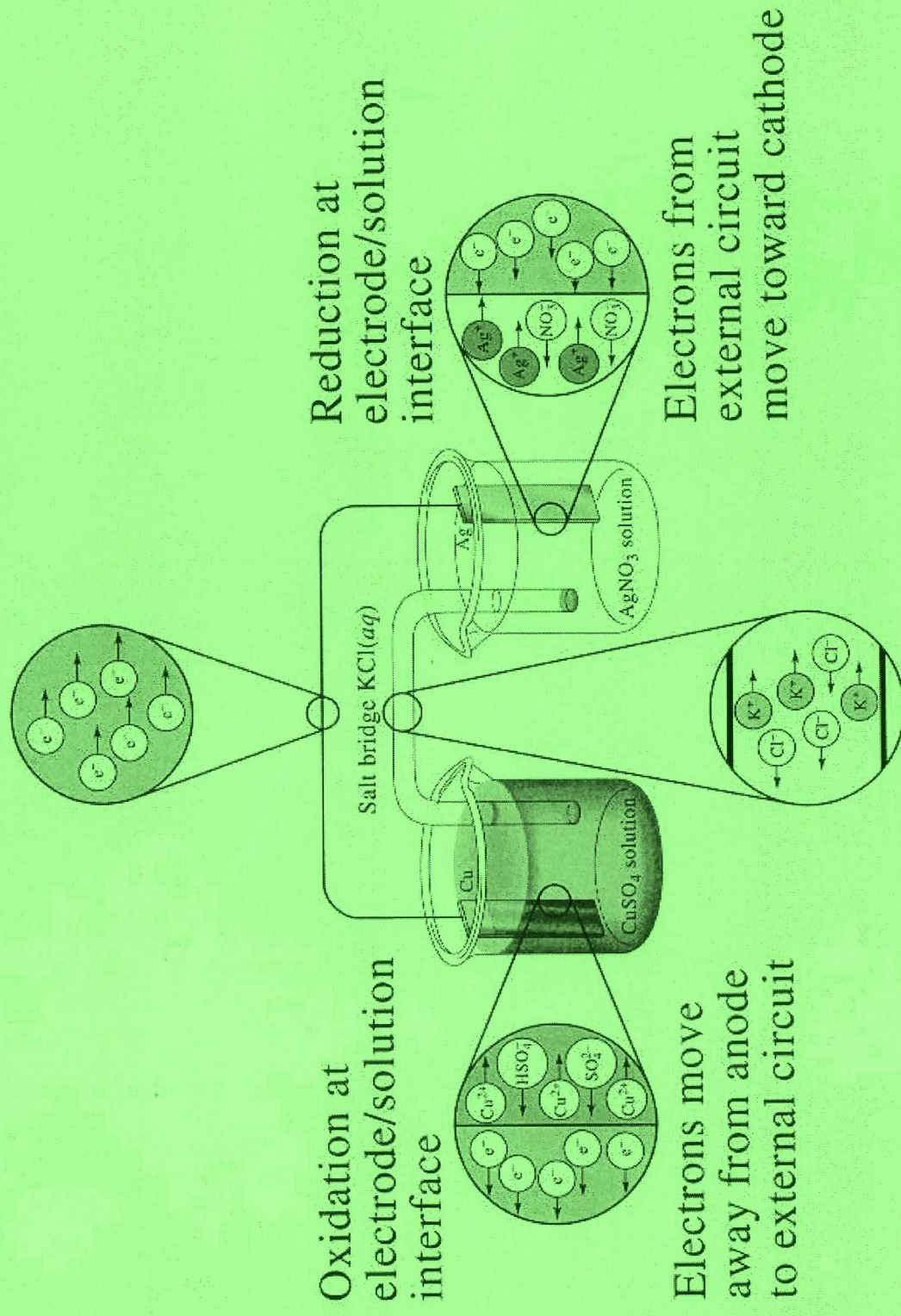
- * The anode is always displayed on the left.
- * The cathode is always displayed on the right.

The vertical line represents the interface.

Current within a cell:-

1. Anions and cations carry electricity within the cell (Ag^+ , Cu^{2+} , K^+ , and the corresponding anions).
2. Current carry electricity within the electrodes and external conductor.
3. Ionic conduction of the electrolyte solution is coupled to the electronic conduction in the electrode by the electrode reactions at the interface.

Currents in Electrochemical Cells



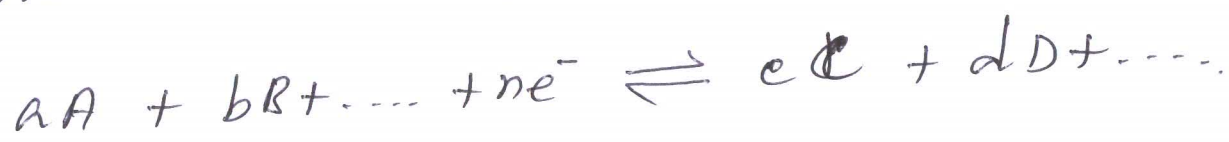
Negative ions in the salt bridge move toward the anode;
positive ions move toward the cathode

Electrode potential: (E)

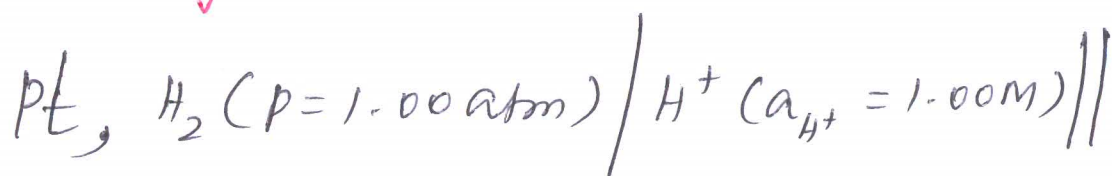
This parameter is a measure of the tendency of gaining or losing electrons.

Standard Electrode Potential (E^0):

Of the half-reaction is defined as its electrode potential when the activities of all reactants and products are unity.

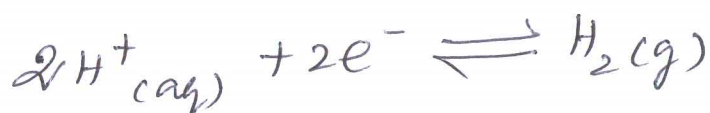


Standard Hydrogen Reference Electrode (SHE):-



* Define this Electrode potential (SHE) = 0

* Electrode potential measurements are comparative (relative) to SHE.



$$E^0_{SHE} = 0.000 \text{ V.}$$

TABLE

I

Standard Electrode Potentials*

Reaction	E^0 at 25°C, V
$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-$	+ 1.359
$\text{O}_2(\text{g}) + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$	+ 1.229
$\text{Br}_2(\text{aq}) + 2\text{e}^- \rightleftharpoons 2\text{Br}^-$	+ 1.087
$\text{Br}_2(\text{l}) + 2\text{e}^- \rightleftharpoons 2\text{Br}^-$	+ 1.065
$\text{Ag}^+ + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+ 0.799
$\text{Fe}^{3+} + \text{e}^- \rightleftharpoons \text{Fe}^{2+}$	+ 0.771
$\text{I}_3^- + 2\text{e}^- \rightleftharpoons 3\text{I}^-$	+ 0.536
$\text{Cu}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+ 0.337
$\text{UO}_2^{2+} + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{U}^{4+} + 2\text{H}_2\text{O}$	+ 0.334
$\text{Hg}_2\text{Cl}_2(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{Hg}(\text{l}) + 2\text{Cl}^-$	+ 0.268
$\text{AgCl}(\text{s}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + \text{Cl}^-$	+ 0.222
$\text{Ag}(\text{S}_2\text{O}_3)_2^{3-} + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + 2\text{S}_2\text{O}_3^{2-}$	+ 0.017
$2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2(\text{g})$	0.000
$\text{AgI}(\text{s}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + \text{I}^-$	- 0.151
$\text{PbSO}_4(\text{s}) + 2\text{e}^- \rightleftharpoons \text{Pb}(\text{s}) + \text{SO}_4^{2-}$	- 0.350
$\text{Cd}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cd}(\text{s})$	- 0.403
$\text{Zn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	- 0.763



* By convention, we write all half-cell reactions as reduction potentials Vs the SHE.

* When written as the reduction potential, we treat the SHE as the anode.

* Thus, when E^0 of a half cell is -ve with respect to the SHE, that electrode is a better anode than the SHE. Hence, SHE acts as Cathode.