
Chemistry Project

Session : 2021-22

ISC

CMS Gomti Nagar Extension

Prepared by:

Sameer Patel

7044549

2225940/019

Acknowledgement

While I was preparing this project file, various information that I found helped me and I am glad that I was able to complete this project and understand many things. Though the preparation of project was an immense learning experience, I inculcated many personal qualities during this process like responsibility, punctuality, confidence and others.

I would like to thank my teachers who supported me all the time, cleared my doubts and to my parents who also played a big role in finalization of my project file. I am taking this opportunity to acknowledge their support and I wish that they keep supporting me like this in the future.

INDEX

Sr no.	Title	Page
1	Electrochemical Cell	1
2	Conductors	2-3
3	Electrolytes	4-5
4	ELECTROLYTIC CONDUCTANCE	6-7
5	KOHLRAUSCH LAW	8-9
6	FARADAY'S LAWS OF ELECTROLYSIS	9
7	Applications of Kohlrausch's law	10-13
8	Electrode Potential	13-15
9	Reference electrode or reference half - cells	16-17
11	EMF of a cell	18-19
12	Nernst Equation	20-22
13	Gibbs Energy relations	23
14	Electrochemical series	24-27

Electrochemistry

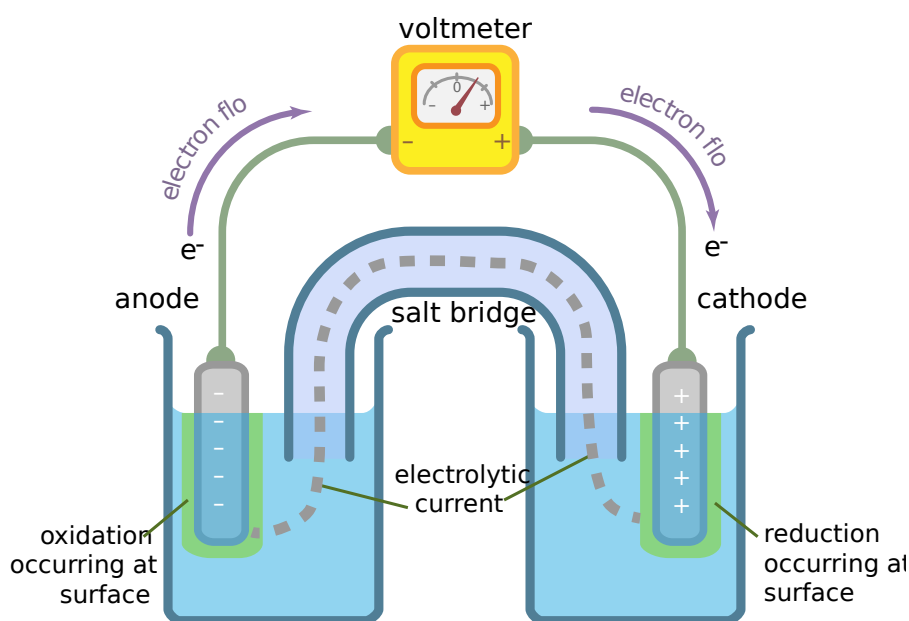
Electrochemistry deals with the conversion of chemical energy into electrical energy and vice-versa.

Electrochemical Cell

An electrochemical cell converts chemical energy into electrical Energy. It involves a redox reaction which results in the decrease of free energy of the system.

When a current is passed through the aqueous solution of an ionic substance, it involves a conversion of electrical energy into chemical energy.

This study of interconversion of chemical energy into electrical and vice – versa extensively studied in a separate branch of chemistry is known as Electrochemistry.

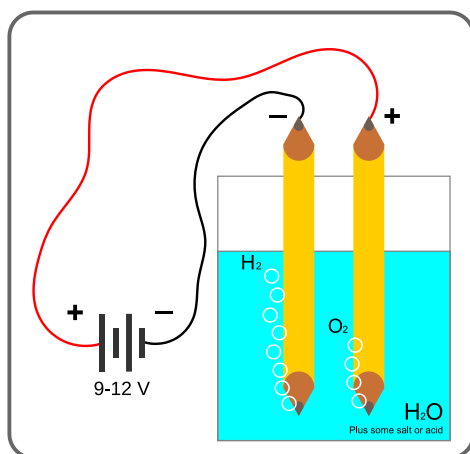
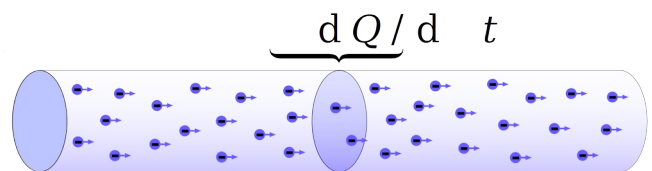


Conductors

A material or an object that conducts electricity is called electric conductor. It helps electric charge (electrons) move from one atom to another easily. Flow of electric charge happens when the voltage is applied across the conductor.

Metallic and Electrolytic conduction

Metallic conduction involves drifting of electrons through vacancies of conduction band in random fashion, brought about under the effect of electric field.



Electrolytic conduction is the movement of free ions in electrolyte which is either the substance in molten state or their aqueous solution.

Electrolysis is an example of electrolytic conduction

Metallic

The flow of electricity takes place without decomposition of the substance.

The conduction of electricity is due to the flow of electrons.

Metallic conductivity decreases with increase in temperature.

Metallic conduction is a physical process that means no new substance is formed.

Electrolytic

The flow of electricity takes place with decomposition of the substance.

The conduction of electricity is due to the flow of ions in the solution.

Electrolytic conduction increases with increases in temperature.

Electrolytic conduction is both physical and chemical change.

Strong Electrolyte

Electrolytes which are completely ionized in aqueous solution or in their molten state, are called strong electrolytes.

Example – all salts, strong acid and strong base

Weak Electrolyte

Electrolytes which are not completely ionized in aqueous solution or in their molten state, are called weak Electrolytes.

Examples :- All carbonic acids (except sulphonic acid), CH_3COOH , HCN , NH_3 , amine, etc.

Electrolysis

The process of decomposition of an electrolyte by the passage of electricity is called electrolysis or electrolytic dissociation.

It is carried out in electrolytic cell where electrical energy is converted into chemical energy. For electrolysis to take place two suitable electrodes are immersed in the liquid.

The solution of an electrolyte contains ions. When an electric potential is applied between the electrodes, the positively charged ions move towards the cathode and negatively ions move towards the anode, when a cation reaches the cathode, it takes up electron(s) and thus gets its charge neutralised. Thus the gain of electrons (decrease in oxidation number) means reduction takes place at the cathode.

Similarly an anion when it reaches the anode, gives up electron(s) and thus gets discharged. Loss of electrons (Increase in oxidation number) means oxidation takes place at anode.

- The tendency of an electrode to lose electrons is known as the oxidation potential.
- The tendency of an electrode to gain electrons is known as the reduction potential.

ORDER OF DISCHARGE POTENTIAL :

Higher be the discharge potential, lower will be the tendency of ion to get discharged at the respective electrode.

The decreasing order of discharge potential or increasing order of deposition of some of ions are given below :

For cations :

K^+ , Na^+ , Ca^{2+} , Mg^{2+} , Al^{3+} , Zn^{2+} , H^+ , Cu^+ , Ag^{2+} , Ag^+

For anions :

SO_4^{4-} , NO_3^{3-} , OH^- , Cl^- , Br^- , I^-

At cathode : $H^+ \xrightarrow{+e} H$

At anode : $Cl^- \xrightarrow{-e} Cl$

ELECTROLYTIC CONDUCTANCE :

Resistance (R) : Metallic and electrolytic conductors obey ohm's law according to which the resistance of a conductor is the ratio of the applied potential difference (V) to the current following(I).

Conductance (C) : The conductance of a conductor is equal to reciprocal of resistance

Specific resistance (ρ) : The resistance (R) of a conductor of uniform cross section is directly proportional to its length and inversely proportional to its area of cross section (A)

$$R = \rho (L/A)$$

Equivalent Conductance (λ_{eq}) :

It is defined as the product of specific conductance (K) and the volume (V in mL) of the solution which contain one gram equivalent of the electrolyte.

$$\lambda_{eq} = K \times V$$

If concentration of solution is C - gram equivalent per litre then
 $\lambda_{eq} = (K \times 1000)/C$

FACTORS INFLUENCING THE CONDUCTIVITY OF ELECTROLYTES :

- Inter ionic attraction :

If inter ionic attraction between solute is more, then the conductivity will be less.

- Polarity of solvent :

If the solvent is greater polarized then the ionization and conductivity will be more.

- Viscosity of medium :

By increasing the viscosity of medium, the conductivity decreases.

- Dilution :

The degree of ionisation increases with the increase of dilution of the solution the conductivity is increases due to increasing the number of ions.

KOHLRAUSCH LAW :

This law states that the equivalent conductivity of any electrolyte at infinite dilution (λ_{eq}^{∞}) is the sum of ionic conductances of the cation and anion given by the electrolytes at infinite dilution.

$$\lambda_{eq}^{\infty} = \lambda_a^{\infty} + \lambda_c^{\infty}$$

Application : This law can be used to

- Determine the equivalent conductivities of weak electrolytes at infinite dilution.
- Determination of degree of dissociation.

FARADAY'S LAWS OF ELECTROLYSIS :

Faraday's first law of electrolysis :

This law states that "The amount of a substance deposited or dissolved at an electrode is directly proportional to the charge passing through the electrolytes.

If a current of I amperes is passed for t seconds, the quantity of charge Q in coulombs. If W , gram of substances is deposited by Q coulombs of electricity, then

$$W \propto Q \propto It$$

$$W = Z It = (E/96500)It$$

Applications of Kohlrausch's law :

Some typical applications of the Kohlrausch's law are described below :

(i) Determination of Λ^∞_m for weak electrolytes :

The molar conductivity of a weak electrolyte at infinite dilution cannot be determined by extrapolation method. However, values for weak electrolytes can be determined by using the Kohlrausch's equation.

(ii) Determination of the degree of ionisation of a weak electrolyte :

The Kohlrausch's law can be used for determining the degree of ionisation of a weak electrolyte at any concentration. If λ_m^c is the molar conductivity of a weak electrolyte at any concentration C and, λ_m^∞ is the molar conductivity of a electrolyte at infinite dilution, then

$$\alpha_c = \frac{\Lambda_m^c}{\Lambda_m^\infty} = \frac{\Lambda_m^c}{(\nu_+ \lambda_+^\infty + \nu_- \lambda_-^\infty)}$$

(iii) Determination of the ionisation constant of a weak electrolyte :

Weak electrolytes in aqueous solutions ionise to a very small extent. The extent of ionisation is described in terms of the degree of ionisation (α).

(iv) Determination of the solubility of a sparingly soluble salt :

The solubility of a sparingly soluble salt in a solvent is quite low. Even a saturated solution of such a salt is so dilute that it can be assumed to be at infinite dilution. Then, the molar conductivity of a sparingly soluble salt at infinite dilution can be obtained from the relationship

Electrode Potential

When a metal (M) is placed in a solution of its ions, either of the following three possibilities can occur, according to the electrode potential solution pressure theory of Nernst.

(i) A metal ion collides with the electrode, and undergoes no change.

(ii) A metal ion collides with the electrode, gains n electrons and gets converted into a metal atom M, (i.e. the metal ion is reduced)

(iii) A metal atom on the electrode M may lose n electrons to the electrode

Thus, “the electrode potential is the tendency of an electrode to lose or gain electrons when it is in contact with solution of its own ions.”

The magnitude of electrode potential depends on the following factors,

- (i) Nature of the electrode,
- (ii) Concentration of the ions in solution,
- (iii) Temperature.

Types of electrode potential :

Depending on the nature of the metal electrode to lose or gain electrons, the electrode potential may be of two types,

Oxidation potential :

When electrode is negatively charged with respect to solution, i.e., it acts as anode. Oxidation occurs.

(ii) Reduction potential : When electrode is positively charged with respect to solution, i.e. it acts as cathode. Reduction occurs.

Standard electrode potential :

“If in the half cell, the metal rod (M) is suspended in a solution of one molar concentration, and the temperature is kept at 298 K, the electrode potential is called standard electrode potential.

The standard electrode potential of a metal may be defined as “the potential difference in volts developed in a cell consisting of two electrodes, the pure metal in contact with a molar solution of one of its ions and the normal hydrogen electrode (NHE)”.

Reference electrode or reference half - cells

It is not possible to measure the absolute value of the single electrode potential directly. Only the difference in potential between two electrodes can be measured experimentally.

It is, therefore, necessary to couple the electrode with another electrode whose potential is known. This electrode is termed as reference electrode or reference half - cells.

Various types of half - cells have been used to make complete cell with spontaneous reaction in forward direction. These half - cells have been summarised in following table

Type	Example	Half - cell reaction	Q =	Reversible to	Electrode Potential (oxdn), E =
Gas ion half - cell	$Pt(H_2) H^+(aq)$ $Pt(Cl_2) Cl^-(aq)$	$\frac{1}{2} H_2(g) \rightarrow H^+(aq) + e^-$ $Cl^-(aq) \rightarrow \frac{1}{2} Cl_2(g) + e^-$	$\frac{[H^+]}{[Cl^-]}$	H^+ Cl^-	$E^0 - 0.0591 \log[H^+]$ $E^0 + 0.0591 \log[Cl^-]$
Metal - metal ion half - cell	$Ag Ag^+(aq)$	$Ag(s) \rightarrow Ag^+(aq) + e^-$	$[Ag^+]$	Ag^+	$E^0 - 0.0591 \log[Ag^+]$
Metal insoluble salt anion half - cell	$Ag, AgCl Cl^-(aq)$	$Ag(s) + Cl^-(aq) \rightarrow AgCl(s) + e^-$	$\frac{1}{[Cl^-]}$	Cl^-	$E^0 + 0.0591 \log[Cl^-]$
Calomel electrode	$Hg, Hg_2Cl_2 Cl^-(aq)$	$2Hg(l) + 2Cl^-(aq) \rightarrow Hg_2Cl_2(s) + 2e^-$	$\frac{1}{[Cl^-]^2}$	Cl^-	$E^0 + 0.0591 \log[Cl^-]$
Metal - metal oxide hydroxide half - cell	$Hg, HgO OH^-(aq)$	$Hg(l) + 2OH^-(aq) \rightarrow HgO(s) + H_2O(l) + 2e^-$	$\frac{1}{[OH^-]^2}$	OH^-	$E^0 + 0.0591 \log[OH^-]$
Oxidation - reduction half - cell	$Pt Fe_{(aq)}^{2+}, Fe_{(aq)}^{3+}$	$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$	$\frac{[Fe^{3+}]}{[Fe^{2+}]}$	Fe^{2+}, Fe^{3+}	$E^0 - 0.0591 \log \frac{[Fe^{3+}]}{[Fe^{2+}]}$

Cell potential or EMF of the cell

(1) "The difference in potentials of the two half - cells of a cell known as electromotive force (emf) of the cell or cell potential."

The difference in potentials of the two half - cells of a cell arises due to the flow of electrons from anode to cathode and flow of current from cathode to anode.

(2) The emf of the cell or cell potential can be calculated from the values of electrode potentials of two the half - cells constituting the cell. The following three methods are in use :

(i) When oxidation potential of anode and reduction potential of cathode are taken into account

(ii) When reduction potentials of both electrodes are taken into account

(iii) When oxidation potentials of both electrodes are taken into account

Emf	Potential difference
It is the potential difference between two electrodes when no current is flowing in the circuit.	It is the difference of the electrode potentials of the two electrodes when the cell is under operation.
It is the maximum voltage that the cell can deliver.	It is always less than the maximum value of voltage which the cell can deliver.

(4) Cell EMF and the spontaneity of the reaction :

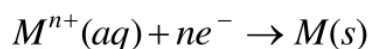
We know, $\Delta G = -nFE_{cell}$

Nature of reaction	ΔG(or ΔG°)	E_{cell} (or E_{cell}°)
Spontaneous	–	+
Equilibrium	0	0
Non – spontaneous	+	–

Nernst's equation

(1) Nernst's equation for electrode potential

The potential of the electrode at which the reaction,



takes place is described by the equation,

$$E_{M^{n+}/M} = E_{M^{n+}/M}^0 - \frac{RT}{nF} \ln \frac{[M(s)]}{[M^{n+}(aq)]}$$

$$\text{or } E_{M^{n+}/M} = E_{M^{n+}/M}^0 - \frac{2.303 RT}{nF} \log \frac{[M(s)]}{[M^{n+}(aq)]}$$

above eq. is called the **Nernst equation**.

Where,

$E_{M^{n+}/M}$ = the potential of the electrode at a given concentration,

$E_{M^{n+}/M}^0$ = the standard electrode potential

R = the universal gas constant, $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$

T = the temperature on the absolute scale,

n = the number of electrons involved in the electrode reaction,

F = the Faraday constant : (96500 C),

$[M(s)]$ = the concentration of the deposited metal,

$[M^{n+}(aq)]$ = the molar concentration of the metal ion in the solution,

The concentration of pure metal $M(s)$ is taken as unity. So, the Nernst equation for the M^{n+}/M electrode is written as,

$$E_{M^{n+}/M} = E_{M^{n+}/M}^0 - \frac{2.303 RT}{nF} \log \frac{1}{[M^{n+}(aq)]}$$

At 298 K, the Nernst equation for the M^{n+}/M electrode can be written as,

$$E_{M^{n+}/M} = E_{M^{n+}/M}^0 - \frac{0.0591}{n} \log \frac{1}{[M^{n+}(aq)]}$$

For an electrode (half - cell) corresponding to the electrode reaction,

Oxidised form + $ne^- \rightarrow$ Reduced form

The Nernst equation for the electrode is written as,

$$E_{\text{half-cell}} = E_{\text{half-cell}}^0 - \frac{2.303 RT}{nF} \log \frac{[\text{Reduced form}]}{[\text{Oxidised form}]}$$

At 298 K, the Nernst equation can be written as,

$$E_{\text{half-cell}} = E_{\text{half-cell}}^0 - \frac{0.0591}{n} \log \frac{[\text{Reduced form}]}{[\text{Oxidised form}]}$$

(2) Nernst's equation for cell EMF

For a cell in which the net cell reaction involving n electrons is, $aA + bB \rightarrow cC + dD$

The Nernst equation is written as,

$$E_{cell} = E_{cell}^0 - \frac{RT}{nF} \ln \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Where, $E_{cell}^0 = E_{cathode}^0 - E_{anode}^0$.

The E_{cell}^0 is called the standard cell potential.

$$\text{or } E_{cell} = E_{cell}^0 - \frac{2.303 RT}{nF} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

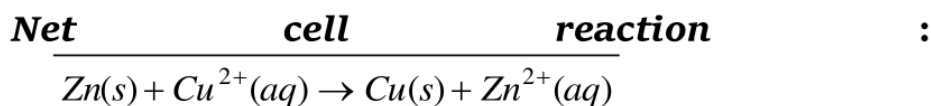
At 298 K, above eq. can be written as,

$$\text{or } E_{cell} = E_{cell}^0 - \frac{0.0592}{n} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

It may be noted here, that the concentrations of A, B, C and D referred in the eqs. are the concentrations at the time the cell *emf* is measured.

(3) Nernst's equation for Daniells cell :

Daniell's cell consists of zinc and copper electrodes. The electrode reactions in Daniell's cell are,



Therefore, the Nernst equation for the Daniell's cell is,

$$E_{cell} = E_{cell}^0 - \frac{2.303 RT}{2F} \log \frac{[Cu(s)][Zn^{2+}(aq)]}{[Zn(s)][Cu^{2+}(aq)]}$$

Since, the activities of pure copper and zinc metals are taken as unity, hence the Nernst equation for the Daniell's cell is,

$$E_{cdll} = E_{cell}^0 - \frac{2.303 RT}{2F} \log \frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]}$$

The above eq. at 298 K is,

$$E_{cdll} = E_{cell}^0 - \frac{0.0591}{2} \log \frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]} V$$

For Daniells cell, $E_{cell}^0 = 1.1 V$

(4) Nernst's equation and equilibrium constant

For a cell, in which the net cell reaction involving n electrons is, $aA + bB \rightarrow cC + dD$

The Nernst equation is

$$E_{Cell} = E_{cell}^0 - \frac{RT}{nF} \ln \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad \text{.....(i)}$$

At equilibrium, the cell cannot perform any useful work. So at equilibrium, E_{Cell} is zero. Also at equilibrium, the ratio

$$\frac{[C]^c [D]^d}{[A]^a [B]^b} = \left[\frac{[C]^c [D]^d}{[A]^a [B]^b} \right]_{equilibrium} = K_c$$

Relationship between potential, Gibbs energy and equilibrium constant

The electrical work (electrical energy) is equal to the product of the *EMF* of the cell and electrical charge that flows through the external circuit i.e.,

$$W_{\max} = nFE_{\text{cell}} \quad \text{.....(i)}$$

According to thermodynamics the free energy change (ΔG) is equal to the maximum work. In the cell work is done on the surroundings by which electrical energy flows through the external circuit, So

$$-W_{\max} = \Delta G$$

$$\text{.....(ii)}$$

from eq. (i) and (ii) $\Delta G = -nFE_{\text{cell}}$

In standard conditions $\Delta G^0 = -nFE_{\text{cell}}^0$

Where ΔG^0 = standard free energy change

$$\text{But } E_{\text{cell}}^0 = \frac{2.303}{nF} RT \log K_c$$

$$\therefore \Delta G^0 = -nF \times \frac{2.303}{nF} RT \log K_c$$

$$\Delta G^0 = -2.303 RT \log K_c \quad \text{or} \quad \Delta G = \Delta G^0 + 2.303 RT \log Q$$

$$\Delta G^0 = -RT \ln K_c \quad (2.303 \log X = \ln X)$$

Electrochemical series

(1) The standard reduction potentials of a large number of electrodes have been measured using standard hydrogen electrode as the reference electrode.

These various electrodes can be arranged in increasing or decreasing order of their reduction potentials. The arrangement of elements in order of increasing reduction potential values is called electrochemical series.

It is also called activity series, of some typical electrodes.

(2) Characteristics of Electrochemical series

(i) The negative sign of standard reduction potential indicates that an electrode when joined with SHE acts as anode and oxidation occurs on this electrode. For example, standard reduction potential of zinc is -0.76 volt, When zinc electrode is joined with SHE, it acts as anode ($-ve$ electrode) i.e., oxidation occurs on this electrode.

Similarly, the $+ve$ sign of standard reduction potential indicates that the electrode when joined with SHE acts as cathode and reduction occurs on this electrode.

(ii) The substances, which are stronger reducing agents than hydrogen are placed above hydrogen in the series and have negative values of standard reduction potentials.

All those substances which have positive values of reduction potentials and placed below hydrogen in the series are weaker reducing agents than hydrogen.

(iii) The substances, which are stronger oxidising agents than H^+ ion are placed below hydrogen in the series.

(iv) The metals on the top (having high negative value of standard reduction potentials) have the tendency to lose electrons readily.

These are active metals. The activity of metals decreases from top to bottom.

The non-metals on the bottom (having high positive values of standard reduction potentials) have the tendency to accept electrons readily. These are active non-metals. The activity of non-metals increases from top to bottom.

Electrochemical Series

Half-Cell Reaction	E^0 (V)
$\text{O}_3 + 2\text{H}^+ + 2e^- = \text{O}_2 + \text{H}_2\text{O}$	+2.08
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2e^- = 2\text{H}_2\text{O}$	+1.78
$\text{N}_2\text{O} + 2\text{H}^+ + 2e^- = \text{N}_2 + \text{H}_2\text{O}$	+1.77
$\text{HClO} + \text{H}^+ + e^- = \frac{1}{2}\text{Cl}_2 + \text{H}_2\text{O}$	+1.63
$2\text{NO} + 2\text{H}^+ + 2e^- = \text{N}_2\text{O} + \text{H}_2\text{O}$	+1.59
$\text{HBrO} + \text{H}^+ + e^- = \frac{1}{2}\text{Br}_{2(\text{aq})} + \text{H}_2\text{O}$	+1.57
$\text{Mn}^{3+} + e^- = \text{Mn}^{2+}$	+1.51
$\text{MnO}_4^- + 8\text{H}^+ + 5e^- = \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.49
$\text{BrO}_3^- + 6\text{H}^+ + 6e^- = \text{Br}^- + 3\text{H}_2\text{O}$	+1.42
$\text{Cl}_{2(\text{g})} + 2e^- = 2\text{Cl}^-$	+1.36
$\text{HCrO}_4^- + 7\text{H}^+ + 3e^- = \text{Cr}^{3+} + 4\text{H}_2\text{O}$	+1.35
$\frac{1}{2}\text{O}_2 + 2\text{H}^+ + 2e^- = \text{H}_2\text{O}$	+1.23
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6e^- = 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	+1.23
$\text{NO}_3^- + 4\text{H}^+ + 3e^- = \text{NO} + 2\text{H}_2\text{O}$	+0.96
$\text{ClO}_{2(\text{aq})} + e^- = \text{ClO}_2^-$	+0.95
$\text{NO}_3^- + 3\text{H}^+ + 2e^- = \text{HNO}_2 + \text{H}_2\text{O}$	+0.93
$\text{Fe}^{3+} + e^- = \text{Fe}^{2+}$	+0.77
$\text{O}_2 + 2\text{H}^+ + 2e^- = \text{H}_2\text{O}_2$	+0.68
$\text{I}_2 + 2e^- = 2\text{I}^-$	+0.54
$\text{H}_2\text{SO}_3 + 4\text{H}^+ + 4e^- = \text{S} + 3\text{H}_2\text{O}$	+0.44
$\text{O}_2 + 2\text{H}_2\text{O} + 4e^- = 4\text{OH}^-$	+0.40
$\text{Fe}_2\text{O}_3 + 4\text{H}^+ + 2e^- = 2\text{FeOH}^+ + \text{H}_2\text{O}$	+0.16
$\text{S} + 2\text{H}^+ + 2e^- = \text{H}_2\text{S}_{(\text{aq})}$	+0.14
$2\text{H}^+ + 2e^- = \text{H}_2$	0.00
$\text{CO}_2 + 2\text{H}^+ + 2e^- = \text{HCOOH}$	-0.20
$\text{Fe}^{2+} + 2e^- = \text{Fe}$	-0.44
$\text{S} + 2e^- = \text{S}^{2-}$	-0.48
$\text{S} + \text{H}_2\text{O} + 2e^- = \text{HS}^- + \text{OH}^-$	-0.48
$\text{Fe}(\text{OH})_3 + e^- = \text{Fe}(\text{OH})_2 + \text{OH}^-$	-0.56

Some Additional Facts :

- ✍ When two or more ions compete at the electrodes then the ion with higher reduction potential gets liberated at the cathode while the one with lower reduction potential at the anode.
- ✍ Cell constant is determined with the help of conductivity bridge, where a standard solution of KCl is used.
- ✍ If the external EMF is slightly more than the actual EMF , the current will flow into the cell and reverse reaction takes place.
- ✍ Identification of cathode and anode is done by the use of galvanometer.
- ✍ KCl / $NaCl$ / NH_4Cl etc., can not be used in the salt bridge of a cell containing silver salt as one of the electrodes as Cl^- ions form a ppt. of $AgCl$ with silver ion.
- ✍ Weston cell is a common example of standard cell. The emf of a standard cell does not change with temperature.
- ✍ In Appolo moon flights, $H_2 - O_2$ fuel cell was the source of energy and drinking water.
- ✍ Conductivity water is the highly purified water whose on conductance is very small. It is prepared by the demineralisation of ordinary water by passing through cation and anion exchange resins.

Bibliography:

Nootan Chemistry

NCERT

Cengage Chemistry