

Revision Notes

Class 12 Chemistry

Chapter 3 – Electrochemistry

Electrochemistry

Electrochemistry is the study of generating electricity from the energy produced during a spontaneous chemical reaction, as well as the application of electrical energy to non-spontaneous chemical changes.

Electrochemical Cells

A spontaneous chemical reaction is one that can occur on its own, and in such a reaction, the system's Gibbs energy falls. This energy is then transformed into electrical energy. It is also feasible to force non-spontaneous processes to occur by providing external energy in the form of electrical energy. Electrochemical Cells are used to carry out these interconversions.

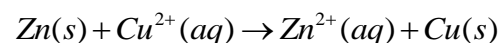
Types

Two types of electrochemical cells are present: **Galvanic cells**, which converts chemical energy into electrical energy and **electrolytic cells** which converts electrical energy into chemical energy.

Galvanic Cells

A spontaneous chemical process or reaction is used to extract cell energy, which is then transformed to electric current.

For example, a Daniell Cell is a Galvanic Cell in which the redox reaction is carried out using Zinc and Copper.



Oxidation Half: $\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^{-}$

Reduction Half: $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$

The reducing agent is Zn , and the oxidising agent is Cu^{2+} . Electrodes are another name for half cells. The anode is the oxidation half, and Cathode is the reduction half. Cathode is a term used to describe a type of electrode. In the external circuit, electrons pass from anode to cathode. Negative polarity is assigned to the anode. Positive polarity is assigned to the cathode. Daniell Cell is a fictional character created by Daniell Cell. The anode is Zn , while the cathode is Cu .

Electrolytic Cell

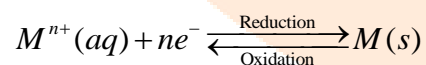
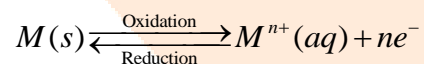
These electrodes are submerged in an electrolytic solution that contains both cations and anions. When current is supplied, the ions migrate towards electrodes of opposite polarity, where they undergo simultaneous reduction and oxidation.

Preferential Discharge of Ions

When more than one cation or anion is present, the discharge process becomes competitive. Any ion that needs to be discharged requires energy, and if there are multiple ions present, the ion that requires the most energy will be discharged first.

Electrode Potential

It can be defined as an element's tendency to lose or gain electrons when in contact with its own ions, causing it to become positively or negatively charged. Depending on whether oxidation or reduction has occurred, the electrode potential will be referred to as oxidation or reduction potential.



Characteristics

- (a) The magnitude and sign of the oxidation and reduction potentials are equal.
- (b) Because E is not a thermodynamic property, its values do not add up.

Standard Electrode Potential (E°)

It can be described as an electrode's electrode potential measured in comparison to a standard hydrogen electrode under standard conditions. The following are the standard conditions:

- (i) A 1M concentration of each ion in the solution.
- (ii) A 298 K temperature.
- (iii) Each gas has a pressure of one bar.

Electrochemical Series

The half-cell potential values are standard and are represented as standard reduction potential values in the table at the conclusion, commonly known as the Electrochemical Series.

Cell Potential or EMF of a Cell

Cell potential is the difference between the electrode potentials of two half cells. If no current is pulled from the cell, it is known as electromotive force (EMF).

$$E_{cell} = E_{cathode} + E_{anode}$$

For this equation we take oxidation potential of anode and reduction potential of cathode.

Since anode is put on left and cathode on right, it follows therefore:

$$= E_R + E_L$$

For a Daniel Cell, therefore:

$$E_{cell}^{\circ} = E_{Cu^{2+}/Cu}^{\circ} - E_{Zn/Zn^{2+}}^{\circ} = 0.34 + (0.76) = 1.10 V$$

Cell Diagram or Representation of a Cell

In accordance with IUPAC recommendations, the following conventions or notations are used to write the cell diagram. The Daniel cell has the following representation:



- (a) The anode half cell is written on the left, while the cathode half cell is written on the right.
- (b) The metal is separated from an aqueous solution of its own ions by a single vertical line.

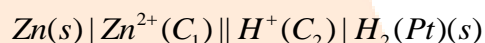
Anodic Chamber: $Zn(s) | Zn^{2+}(aq)$

Cathodic Chamber: $Cu^{2+}(aq) | Cu(s)$

- (c) A salt bridge is represented by a double vertical line.
- (d) After the formula of the corresponding ion, the molar concentration (C) is placed in brackets.
- (e) The cell's e.m.f. value is written on the cell's extreme right side. As an example:

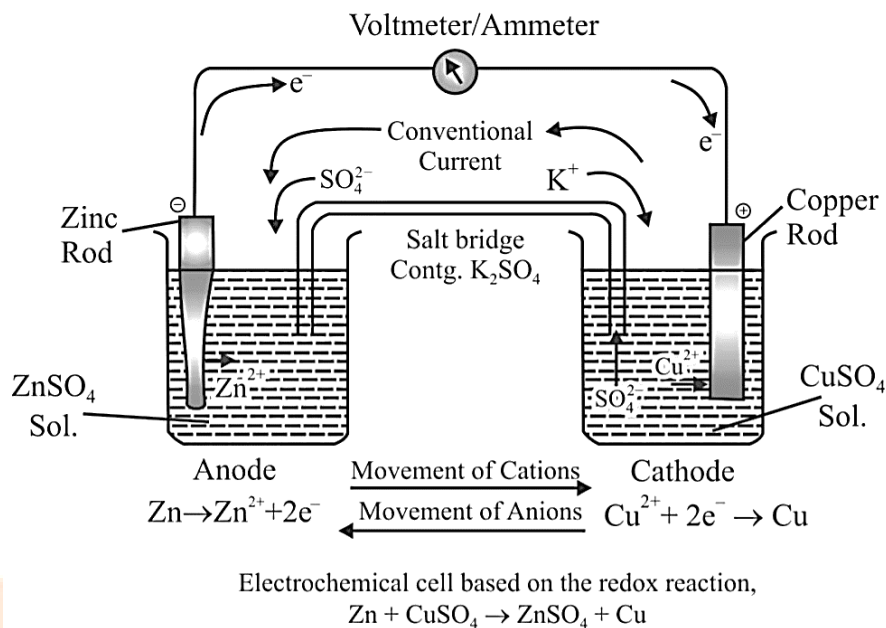


- (f) If an inert electrode, such as platinum, is used in the cell's construction, it may be written in brackets alongside the working electrode, as when a zinc anode is coupled to a hydrogen electrode.



Salt Bridge

The salt bridge maintains charge balance and completes the circuit by allowing ions to flow freely through it. It contains a gel containing an inert electrolyte such as Na_2SO_4 or KNO_3 . Through the salt bridge, negative ions travel to the anode and positive ions flow to the cathode, maintaining charge balance and allowing the cell to function.



Spontaneity of a Reaction

$$\Delta G = -nFE_{\text{cell}}$$

ΔG should be negative and cell potential should be positive for a spontaneous cell reaction.

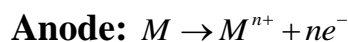
In the following equation, if we take the standard value of cell potential, we will also get the standard value of ΔG .

$$\Delta G^\circ = -nFE_{\text{CELL}}^\circ$$

Types of Electrodes

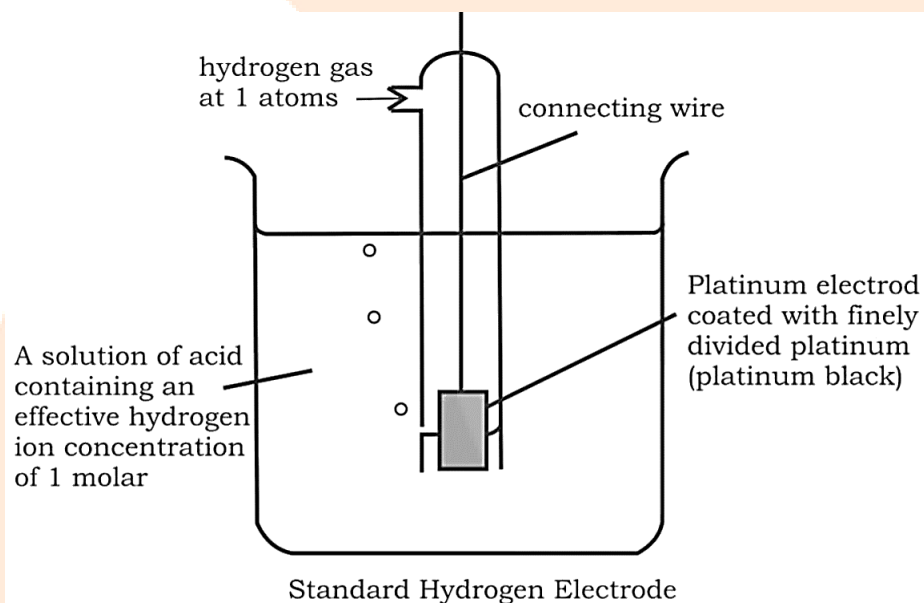
Metal – Metal Ion electrodes

An electrolyte solution containing metal ions is dipped into a metal rod/plate. Because of the potential difference between these two phases, this electrode can function as both a cathode and an anode.



Gas Electrodes

Electrode gases such as H_2 and Cl_2 are used in conjunction with their respective ions. H_2 gas, for example, is utilised in conjunction with a dilute solution of HCl (H^+ ions). To avoid reacting with the acid, the metal should be inert.



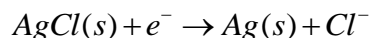
Anode: $H_2 \rightarrow 2H^+ + 2e^-$

Cathode: $2H^+ + 2e^- \rightarrow H_2$

The hydrogen electrode is also used as a standard for measuring the potentials of other electrodes. As a reference, its own potential is set at $0V$. The concentration of the HCl used as a reference is $1M$, and the electrode is known as the "Standard Hydrogen Electrode (SHE)".

Metal – Insoluble Salt Electrode

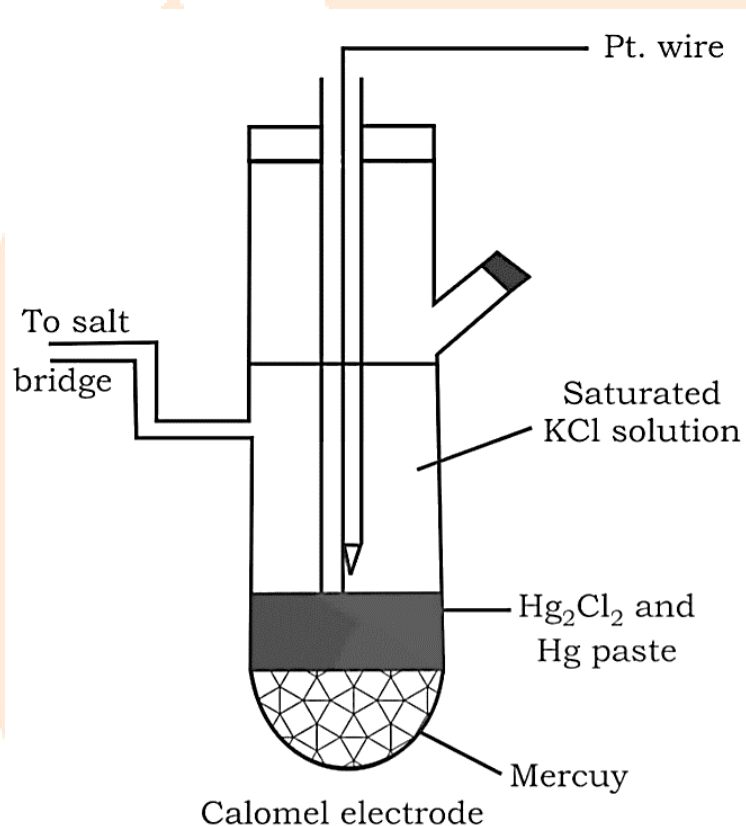
As electrodes, we use salts of several metals that are only sparingly soluble with the metal itself. When we employ $AgCl$ with Ag , for example, there is a potential gap between these two phases, as seen in the following reaction:



This electrode is made by dipping a silver rod in a solution containing $AgCl(s)$ and Cl^- ions.

Calomel Electrode

Mercury is combined with two other phases: calomel paste (Hg_2Cl_2) and a Cl^- ions containing electrolyte.



Cathode: $Hg_2Cl_2(s) + 2e^- \rightarrow 2Hg(l) + 2Cl^-(aq)$

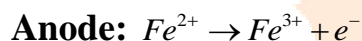
Anode: $2Hg(l) + 2Cl^-(aq) \rightarrow Hg_2Cl_2(s) + 2e^-$

This electrode is also utilised as a reference point for determining other potentials. It's also known as Standard Calomel Electrode in its standard form (SCE).

Redox Electrode

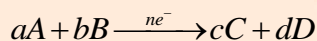
Two distinct oxidation states of the same metal are used in the same half cell in these electrodes. For example, Fe^{2+} and Fe^{3+} are dissolved in the same container and the electron transfer is performed using a platinum inert electrode.

The following reactions may occur:



Nernst Equation

It establishes a link between electrode voltage and ion concentration. When a result, as the concentration of ions rises, so does the reduction potential. For a type of generic electrochemical reaction.



Nernst equation can be given as:

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

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{2303}{nF} RT \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Substituting the values of R and F we get:

$$E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.0591}{n} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Applications of Nernst Equation

Equilibrium Constant from Nernst Equation

For a Daniel Cell, at equilibrium

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

$$E_{\text{cell}}^0 = \frac{2.303RT}{2F} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

But at equilibrium:

$$\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = K_c$$

$$E_{\text{cell}}^0 = \frac{2.303RT}{2F} \log K_c$$

$$\begin{aligned} E_{\text{cell}}^0 &= \frac{2.303 \times 8.314 \times 298}{2 \times 96500} \log K_c \\ &= \frac{0.0591}{2} \log K_c \end{aligned}$$

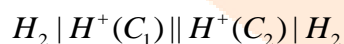
In general:

$$E_{\text{cell}}^0 = \frac{0.0591}{n} \log K_c$$

$$\log K_c = \frac{nE_{\text{cell}}^0}{0.0591}$$

Concentration Cells

Concentration cells are formed when two electrodes of the same metal are dipped individually into two solutions of the same electrolyte with varying concentrations and the solutions are connected by a salt bridge. As an example:



These are of two types:

Electrode Concentration Cells

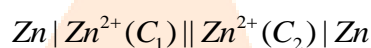


$$E_{\text{cell}} = 0 - \frac{0.059}{n} \log \frac{P_2}{P_1}$$

Where, $P_2 < P_1$ for spontaneous reaction.

Electrolyte Concentration Cell

The EMF of concentration cell at 298 K is given by:

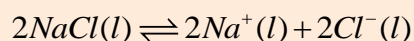


$$E_{\text{cell}} = \frac{0.0591}{n_1} \log \frac{c_2}{c_1}$$

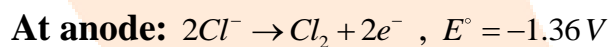
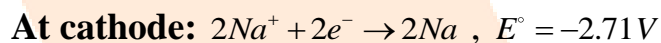
Where, $C_2 > C_1$ for spontaneous reaction

Cases of Electrolysis

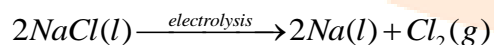
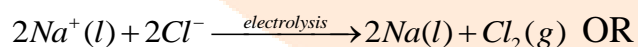
Electrolysis of Molten Sodium Chloride



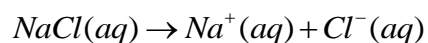
The reactions occurring at the two electrodes may be shown as follows:

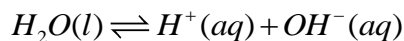


Overall reaction:



Electrolysis of an aqueous solution of Sodium Chloride



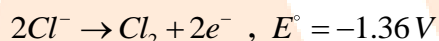
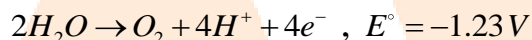


At cathode:



Thus H_2 gas is evolved at cathode value Na^+ ions remain in solution.

At anode:



Thus, Cl_2 gas is evolved at the anode by over voltage concept while OH^- ions remain in the solution.

Batteries

The term "battery" refers to a configuration in which Galvanic cells are connected in series to achieve a higher voltage.

Primary Batteries

Primary cells can be employed indefinitely as long as active components are present. When they're gone, the cell stops working and can't be used again. For instance, a Dry Cell or a Leclanche Cell, as well as a Mercury Cell.

Dry Cell

Anode: Zinc container

Cathode: Carbon (graphite) rod surrounded by powdered MnO_2 and carbon

Electrolyte: NH_4Cl and $ZnCl_2$

Reaction:

Anode: $Zn \rightarrow Zn^{2+} + 2e^{-}$

Cathode: $MnO_2 + NH_4^{+} + e^{-} \rightarrow MnO(OH) + NH_3$

The standard potential of this cell is 1.5 V, which decreases as the battery is repeatedly discharged, and it cannot be refilled once used.

Mercury Cells

These are used in small equipments like watches, hearing aids.

Anode: $Zn-Hg$ Amalgam

Cathode: Paste of HgO and carbon

Electrolyte: Paste of KOH and ZnO

Anode: $Zn(Hg) + 2OH^{-} \rightarrow ZnO(s) + H_2O + 2e^{-}$

Cathode: $HgO(s) + H_2O + 2e^{-} \rightarrow Hg(l) + 2OH^{-}$

Overall Reaction: $Zn(Hg) + HgO(s) \rightarrow ZnO(s) + Hg(l)$

The cell potential is approximately 1.35 V and remains constant during its life.

Secondary Batteries

Secondary batteries are rechargeable for many applications and can be recharged multiple times. Lead storage batteries and $Ni-Cd$ batteries, for example.

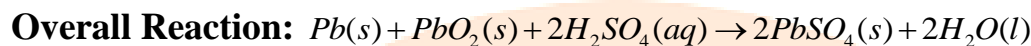
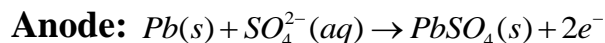
Lead Storage Battery

Anode: Lead (Pb)

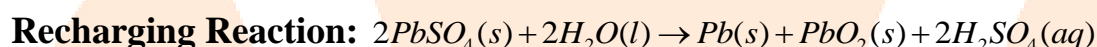
Cathode: Grid of lead packed with lead oxide (PbO_2)

Electrolyte: 38% solution of H_2SO_4

Discharging Reaction

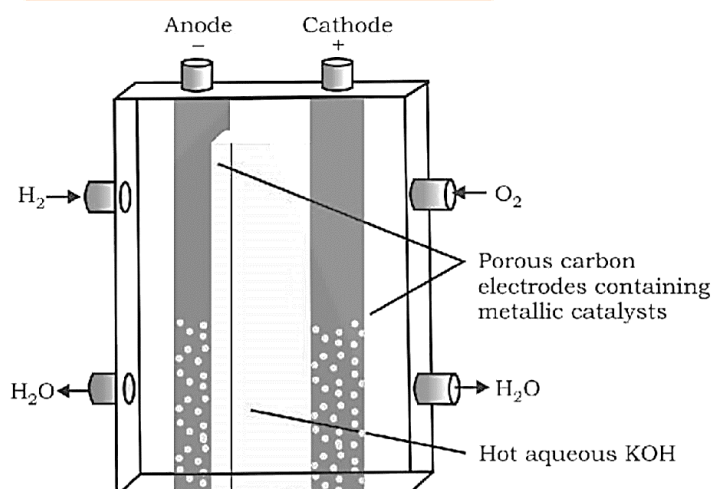
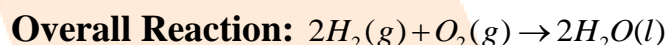


To recharge the cell, it is connected to a higher-potential cell, which acts as an electrolytic cell and reverses the processes. At the relevant electrodes, $Pb(s)$ and $PbO_2(s)$ are regenerated. These cells produce a voltage that is nearly constant.



Fuel Cells

A fuel cell varies from a traditional battery in that the reactants are supplied externally from a reservoir rather than being stored inside the cell. In space vehicles, fuel cells are employed, and the two gases are supplied from external storage. The electrodes in this cell are carbon rods, and the electrolyte is KOH .



Corrosion

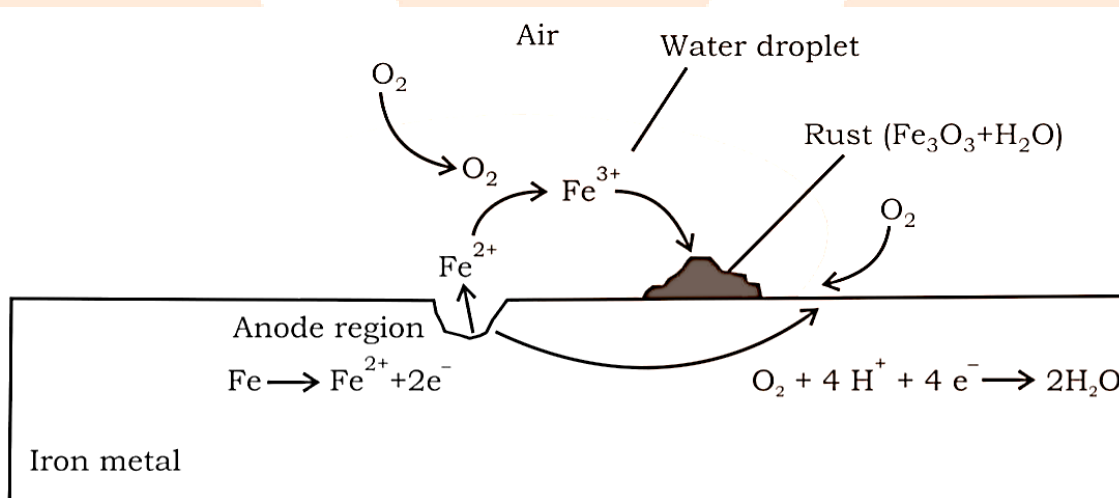
On the surface of iron or any other metal, it entails a redox process and the development of an electrochemical cell.

The oxidation of iron (anode) occurs at one point, while the reduction of oxygen to generate water occurs at another (cathode). Fe is first oxidised to Fe^{2+} , which is then converted to Fe^{3+} in the presence of oxygen, which subsequently combines with water to generate rust, which is represented by $Fe_2O_3 \cdot xH_2O$.

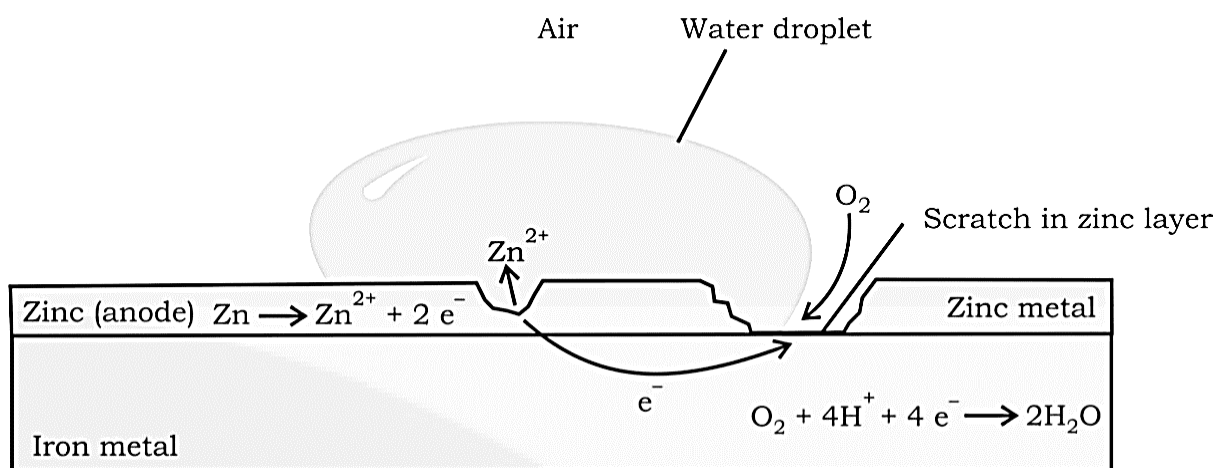
Anode: $2Fe(s) \rightarrow 2Fe^{2+} + 4e^-$, $E^\circ = +0.44 V$

Cathode: $O_2(g) + 4H^+ + 4e^- \rightarrow 2H_2O(l)$, $E^\circ = 1.23 V$

Overall Reaction: $2Fe(s) + O_2(g) + 4H^+ \rightarrow 2Fe^{2+} + 2H_2O$, $E^\circ_{Cell} = 1.67 M$



Painting or coating iron with other metals, such as zinc, helps prevent it from rusting. Galvanisation is the name for the latter procedure. Because Zn has a higher potential to oxidise than iron, it is oxidised first, while iron is protected. Cathodic Protection is another name for this approach of shielding one metal by the other.



Conductance (G)

It is defined as the ease with which electric current passes through a conductor and is the reciprocal of resistance.

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

SI unit is Siemen (S).

$$1\text{ S} = 1\text{ ohm}^{-1}(\text{mho})$$

Conductivity

It is the reciprocal of resistivity (ρ) .

$$\kappa = \frac{1}{\rho} = \frac{1}{R} \times \frac{\ell}{A} = G \times \frac{\ell}{A}$$

Now is $\ell = 1\text{ cm}$ and $A = 1\text{ cm}^2$, then $\kappa = G$

As a result, the conductivity of an electrolytic solution can be defined as the conductance of a 1 cm long solution with a 1 cm^2 cross-sectional area.

Factors Affecting Electrolyte Conductance

Electrolyte

In a dissolved or molten form, an electrolyte is a substance that dissociates in solution to produce ions and hence conducts electricity.

Examples: HCl , $NaOH$, KCl are strong electrolytes and CH_3COOH , NH_4OH are weak electrolytes.

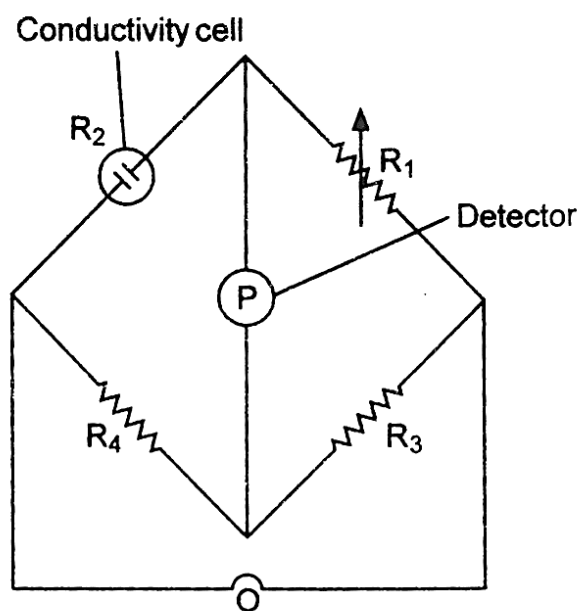
Electrolytic or ionic conductance refers to the conductance of electricity by ions present in solutions. The flow of electricity through an electrolyte solution is governed by the following factors.

- (i) **Electrolyte nature or interionic attractions:** The lower the solute-solute interactions, the larger the freedom of ion mobility and the higher the conductance.
- (ii) **Ion Solvation:** As the amount of solute-solvent interactions increases, the extent of solvation increases, and the electrical conductance decreases.
- (iii) **The nature of the solvent and its viscosity:** The larger the solvent-solvent interactions, the higher the viscosity, and the greater the solvent's resistance to ion flow, and thus the lower the electrical conductance.
- (iv) **Temperature:** As the temperature of an electrolytic solution rises, solute-solute, solute-solvent, and solvent-solvent interactions diminish, causing electrolytic conductance to rise.

Measurement of Conductance

As we know, $\kappa = \frac{1}{R} \times \frac{\ell}{A}$

If we measure ℓ , A , and R , we can figure out what the value of κ is. Using the 'Wheatstones' bridge method, the resistance of the solution R between two parallel electrodes is calculated.



Wheatstone Bridge

It is made up of two fixed resistances, R_3 and R_4 , a variable resistance R_1 , and a conductivity cell with an unknown resistance, R_2 . When no current goes through the detector, the bridge is balanced. Then, in these circumstances:

$$\frac{R_1}{R_2} = \frac{R_3}{R_4} \quad \text{or} \quad R_2 = \frac{R_1 R_4}{R_3}$$

Molar Conductivity

It's the total conducting power of all the ions created by dissolving one mole of an electrolyte between two big electrodes separated by one centimetre.

Mathematically,

$$\Lambda_m = \kappa \times V, \quad \Lambda_m = \frac{\kappa \times 1000}{C}$$

where, V is the volume of solution in cm^3 containing 1 mole of electrolyte and C is the molar concentration.

Units: $\Lambda_m = \frac{\kappa \times 1000}{C} = \frac{\text{Scm}^{-1}}{\text{molcm}^{-3}}$
 $= \text{ohm}^{-1}\text{cm}^2\text{mol}^{-1}$ or $\text{Scm}^2\text{mol}^{-1}$

Equivalent Conductivity

It is the electrical conductivity of one equivalent of electrolyte placed between two big electrodes separated by one centimetre.

Mathematically:

$$\Lambda_{\text{eq}} = \kappa \times v =$$

$$\Lambda_{\text{eq}} = \frac{\kappa \times 1000}{N}$$

Where, v is the volume of solution in cm^3 containing 1 equivalent of electrolyte and N is normality.

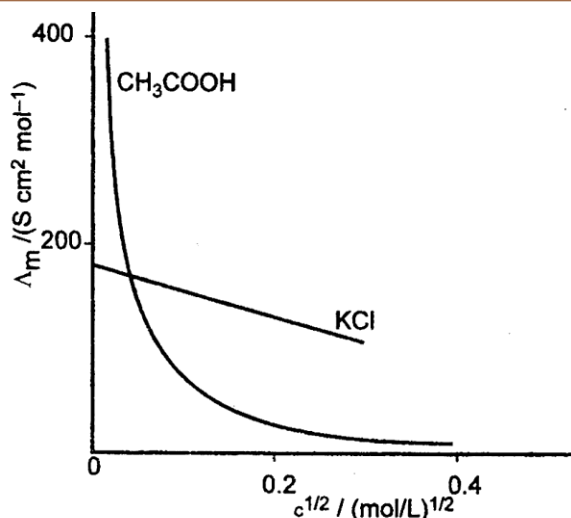
Units: $\Lambda_{\text{eq}} = \frac{\kappa \times 1000}{N}$

$$= \frac{\text{Scm}^{-1}}{\text{equivalent cm}^{-3}} = \frac{\text{Ohm}^{-1}\text{cm}^2 \text{ equivalent}^{-1}}{\text{Scm}^2 \text{ equivalent}^{-1}}$$

Variation of Conductivity and Molar Conductivity with Dilution

Because the number of ions per unit volume that carry the current in the solution reduces as concentration lowers, conductivity drops. With a decrease in concentration, molar conductivity rises. This is due to an increase in the total volume V of a solution containing one mole of electrolyte. The drop in κ as a result of dilution of a solution has been found to be more than compensated by increases in its volume.

Graphical representation of the variation of Λ_m vs \sqrt{c} .



Limiting Molar Conductivity (Λ_m)

Limiting molar conductivity, also known as molar conductivity at infinite dilution, is the value of molar conductivity as the concentration approaches zero. In the case of a strong electrolyte, extrapolation of the Λ_m vs \sqrt{c} curve can be used to derive the molar conductivity at infinite dilution. Extrapolation of the curve, on the other hand, cannot be used to calculate the value of molar conductivity of a weak electrolyte at infinite dilution since the curve becomes practically parallel to the y-axis as concentration approaches zero.

The mathematical relationship between Λ_m and Λ_m° for strong electrolyte was developed by Debye, Huckel and Onsagar. In simplified form the equation can be given as:

$$\Lambda_m = \Lambda_m^\circ - bc^{1/2}$$

Kohlrausch's Law

It asserts that an electrolyte's limiting molar conductivity can be described as the total of the individual contributions of the electrolyte's anion and cation. In general, if an electrolyte produces v_+ cations and v_- anions upon dissociation, its limiting molar conductivity is given by:

$$\Lambda_m^\circ = v_+ \lambda_+^\circ + v_- \lambda_-^\circ$$

Applications of Kohlrausch's Law

(i) Calculation of molar conductivities of weak electrolyte at infinite dilution

For example, the molar conductivity of acetic acid at infinite dilution can be calculated using the molar conductivities of strong electrolytes like HCl , CH_3COONa , and $NaCl$ at infinite dilution, as shown below.

$$\Lambda_{m(CH_3-COOH)}^{\circ} = \Lambda_{m(CH_3-COO Na)}^{\circ} + \Lambda_{m(HCl)}^{\circ} - \Lambda_{m(NaCl)}^{\circ}$$

(ii) Determination of degree of dissociation of weak electrolytes

$$\text{Degree of dissociation } \alpha = \frac{\Lambda_m^c}{\Lambda_m^{\circ}}$$

(iii) Determination of dissociation constant of weak electrolytes:

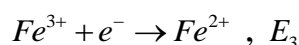
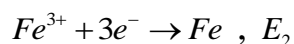
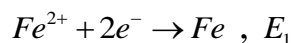
$$K = \frac{c\alpha^2}{1-\alpha}$$

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^{\infty}}$$

$$K = \frac{c(\Lambda_m^c / \Lambda_m^{\infty})^2}{1 - \Lambda_m^c / \Lambda_m^{\infty}} = \frac{C(\Lambda_m^c)^2}{\Lambda_m^{\infty}(\Lambda_m^{\infty} - \Lambda_m^c)}$$

Use of ΔG in Relating EMF values of Half Cell Reactions

When we have two half-cell reactions that produce another half-cell reaction when we combine them, their emfs cannot be mixed directly. However, thermodynamic functions such as ΔG can be added and EMF values can be connected through them in any scenario. Take a look at the three half-cell responses below:



We can clearly see that subtracting the first reaction from the second yields the third reaction. However, the same relationship does not hold true for EMF values.

That is: $E_3 \neq E_2 - E_1$. But the ΔG values can be related according to the reactions:

$$\begin{aligned}\Delta G_3 &= \Delta G_2 - \Delta G_1 \\ -n_3 FE_3 &= -n_2 FE_2 + n_1 FE_1 \\ -E_3 &= -3E_2 + 2E_1 \\ \Rightarrow E_3 &= 3E_2 - 2E_1\end{aligned}$$

Formula:

$$1. \quad R = \rho \left(\frac{\ell}{A} \right) = \rho \times \text{Cell constant}$$

Where, R = Resistance,

A = Area of cross-section of the electrodes

ρ = Resistivity

$$2. \quad \kappa = \frac{1}{R} \times \text{cell constant}$$

Where, κ = Conductivity or specific conductance

$$3. \quad \Lambda_m = \frac{\kappa \times 1000}{M}$$

Where, Λ_m = Molar conductivity

M = Molarity of the solution.

$$4. \quad \Lambda_m^\infty (A_x B_y) = x \Lambda_m^\infty (A^y) + y \Lambda_m^\infty (B^{x-})$$

$$5. \quad \alpha = \frac{\Lambda_m^c}{\Lambda_m^\circ}$$

Where, α = Degree of dissociation

Λ_m^c = Molar conductivity at a given concentration

6. For a weak binary electrolyte AB

$$K = \frac{c\alpha^2}{1-\alpha} = \frac{c(\Lambda_m^c)^2}{\Lambda_m^\infty(\Lambda_m^\infty - \Lambda_m^c)}$$

Where, K is the Dissociation constant

$$\begin{aligned} E_{\text{edl}}^\circ &= E_{\text{cathode}}^\circ + E_{\text{anode}}^\circ \\ &= E^\circ_{\text{Right}} + E^\circ_{\text{left}} \end{aligned}$$

7. Nernst equation for a generation electrochemical reaction

$$E_{\text{offl}} = E_{\text{cell}}^\circ - \frac{0.059}{n} \log \frac{[A]^2[B]^b}{[C]^c[D]^d}$$

$$8. \log K_c = \frac{n}{0.0591} E_{\text{cell}}^\circ$$

Where, K_c = Equilibrium constant.

$$9. \Delta_r G^\circ = -nFE_{\text{cell}}^\circ$$

$$\Delta_r G^\circ = -2.303RT \log K_c$$

$\Delta_r G^\circ$ = Standard Gibbs energy of a reaction

$$10. Q = I \times t$$

Where, Q = Quantity of charge in coulombs

I = Current in amperes

t = Time in seconds

$$11. m = Z \times I \times t$$

Where, m = mass of the substance liberated at the electrodes

Z = Electrochemical equivalent

STANDARD REDUCTION POTENTIAL AT 298 K. IN ELECTROCHEMICAL ORDER

$\text{H}_4\text{XeO}_6 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{XeO}_3 + 3\text{H}_2\text{O}$	+3.0	$\text{Hg}_2\text{SO}_4 + 2\text{e}^- \rightarrow 2\text{Hg} + \text{SO}_4^{2-}$	+0.62
$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	+2.87	$\text{MnO}_4^{2-} + 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{MnO}_2 + 4\text{OH}^-$	+0.60
$\text{O}_3 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{O}_2 + \text{H}_2\text{O}$	+2.07	$\text{MnO}_4^- + \text{e}^- \rightarrow \text{MnO}_4^{2-}$	+0.56
$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}$	+2.05	$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	+0.54
$\text{Ag}^{2+} + \text{e}^- \rightarrow \text{Ag}^+$	+1.98	$\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$	+0.52
$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	+1.81	$\text{I}_3^- + 2\text{e}^- \rightarrow 3\text{I}^-$	+0.53
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.78	$\text{NiOOH} + \text{H}_2\text{O} + \text{e}^- \rightarrow \text{Ni(OH)}_2 + \text{OH}^-$	+0.49
$\text{Au}^+ + \text{e}^- \rightarrow \text{Au}$	+1.69	$\text{Ag}_2\text{CrO}_4 + 2\text{e}^- \rightarrow 2\text{Ag} + \text{CrO}_4^{2-}$	+0.45
$\text{Pb}^{4+} + 2\text{e}^- \rightarrow \text{Pb}^{2+}$	+1.67	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	+0.40
$2\text{HClO} + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{Cl}_2 + 2\text{H}_2\text{O}$	+1.63	$\text{ClO}_4^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{ClO}_3^- + 2\text{OH}^-$	+0.36
$\text{Ce}^{4+} + \text{e}^- \rightarrow \text{Ce}^{3+}$	+1.61	$[\text{Fe(CN)}_6]^{3-} + \text{e}^- \rightarrow [\text{Fe(CN)}_6]^{4-}$	+0.36
$2\text{HBrO} + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{Br}_2 + 2\text{H}_2\text{O}$	+1.60	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.34
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.51	$\text{Hg}_2\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Hg} + 2\text{Cl}^-$	+0.27
$\text{Mn}^{3+} + \text{e}^- \rightarrow \text{Mn}^{2+}$	+1.51	$\text{AgCl} + \text{e}^- \rightarrow \text{Ag} + \text{Cl}^-$	+0.22
$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	+1.40	$\text{Bi} + 3\text{e}^- \rightarrow \text{Bi}$	+0.20
$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	+1.36	$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	+0.16
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	+1.33	$\text{Sn}^{4+} + 2\text{e}^- \rightarrow \text{Sn}^{2+}$	+0.15
$\text{O}_3 + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{O}_2 + 2\text{OH}^-$	+1.24	$\text{AgBr} + \text{e}^- \rightarrow \text{Ag} + \text{Br}^-$	+0.07
$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.23		
$\text{ClO}_4^- + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{ClO}_3^- + \text{H}_2\text{O}$	+1.23	$\text{Ti}^{4+} + \text{e}^- \rightarrow \text{Ti}^{3+}$	0.00
$\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}$	+1.23	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0, by definition
$\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^-$	+1.09	$\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}$	-0.04
$\text{Pu}^{4+} + \text{e}^- \rightarrow \text{Pu}^{3+}$	+0.97	$\text{O}_2 + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{HO}_2^- + \text{OH}^-$	-0.08
$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	+0.96	$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.13
$2\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	+0.92	$\text{In}^+ + \text{e}^- \rightarrow \text{In}$	-0.14
$\text{ClO}^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Cl}^- + 2\text{OH}^-$	+0.89	$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.14
$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow \text{Hg}$	+0.86	$\text{AgI} + \text{e}^- \rightarrow \text{Ag} + \text{I}^-$	-0.15
$\text{NO}_3^- + 2\text{H}^+ + \text{e}^- \rightarrow \text{NO}_2 + \text{H}_2\text{O}$	+0.80	$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.23
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.80	$\text{Co}^{2+} + 2\text{e}^- \rightarrow \text{Co}$	-0.28
$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	+0.79	$\text{In}^{3+} + 3\text{e}^- \rightarrow \text{In}$	-0.34
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	+0.77	$\text{Tl}^+ + \text{e}^- \rightarrow \text{Tl}$	-0.34
$\text{BrO}^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Br}^- + 2\text{OH}^-$	+0.76	$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.36

$\text{Ti}^{3+} + \text{e}^- \rightarrow \text{Ti}^{2+}$	-0.37	$\text{V}^{2+} + 2\text{e}^- \rightarrow \text{V}$	-1.19
$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	-0.40	$\text{Ti}^{2+} + 2\text{e}^- \rightarrow \text{Ti}$	-1.63
$\text{In}^{2+} + \text{e}^- \rightarrow \text{In}^+$	-0.40	$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.66
$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$	-0.41	$\text{U}^{3+} + 3\text{e}^- \rightarrow \text{U}$	-1.79
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.44	$\text{Sc}^{3+} + 3\text{e}^- \rightarrow \text{Sc}$	-2.09
$\text{In}^{3+} + 2\text{e}^- \rightarrow \text{In}^+$	-0.44	$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.36
$\text{S} + 2\text{e}^- \rightarrow \text{S}^{2-}$	-0.48	$\text{Ce}^{3+} + 3\text{e}^- \rightarrow \text{Ce}$	-2.48
$\text{In}^{3+} + \text{e}^- \rightarrow \text{In}^{2+}$	-0.49	$\text{La}^{3+} + 3\text{e}^- \rightarrow \text{La}$	-2.52
$\text{U}^{4+} + \text{e}^- \rightarrow \text{U}^{3+}$	-0.61	$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71
$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.74	$\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$	-2.87
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76	$\text{Sr}^{2+} + 2\text{e}^- \rightarrow \text{Sr}$	-2.89
$\text{Cd}(\text{OH})_2 + 2\text{e}^- \rightarrow \text{Cd} + 2\text{OH}^-$	-0.81	$\text{Ba}^{2+} + 2\text{e}^- \rightarrow \text{Ba}$	-2.91
$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.83	$\text{Ra}^{2+} + 2\text{e}^- \rightarrow \text{Ra}$	-2.92
$\text{Cr}^{2+} + 2\text{e}^- \rightarrow \text{Cr}$	-0.91	$\text{Cs}^+ + \text{e}^- \rightarrow \text{Cs}$	-2.92
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.18	$\text{Rb}^+ + \text{e}^- \rightarrow \text{Rb}$	-2.93
		$\text{K}^+ + \text{e}^- \rightarrow \text{K}$	-2.93
		$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.05

Reduction Potential in Alphabetical order:

$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.80	$\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$	-2.87
$\text{Ag}^{2+} + \text{e}^- \rightarrow \text{Ag}^+$	+1.98	$\text{Cd}(\text{OH})_2 + 2\text{e}^- \rightarrow \text{Cd} + 2\text{OH}^-$	-0.81
$\text{AgBr} + \text{e}^- \rightarrow \text{Ag} + \text{Br}^-$	+0.0713	$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	-0.40
$\text{AgCl} + \text{e}^- \rightarrow \text{Ag} + \text{Cl}^-$	+0.22	$\text{Ce}^{3+} + 3\text{e}^- \rightarrow \text{Ce}$	-2.48
$\text{Ag}_2\text{CrO}_4 + 2\text{e}^- \rightarrow 2\text{Ag} + \text{CrO}_4^{2-}$	+0.45	$\text{Ce}^{4+} + \text{e}^- \rightarrow \text{Ce}^{3+}$	+1.61
$\text{AgF} + \text{e}^- \rightarrow \text{Ag} + \text{F}^-$	+0.78	$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	+1.36
$\text{AgI} + \text{e}^- \rightarrow \text{Ag} + \text{I}^-$	-0.15	$\text{ClO}^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Cl}^- + 2\text{OH}^-$	+0.89
$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.66	$\text{ClO}_4^- + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{ClO}_3^- + \text{H}_2\text{O}$	+1.23
$\text{Au}^+ + \text{e}^- \rightarrow \text{Au}$	+1.69	$\text{ClO}_4^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{ClO}_3^- + 2\text{OH}^-$	+0.36
$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	+1.40	$\text{Co}^{2+} + 2\text{e}^- \rightarrow \text{Co}$	-0.28
$\text{Ba}^{2+} + 2\text{e}^- \rightarrow \text{Ba}$	+2.91	$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	+1.81
$\text{Be}^{2+} + 2\text{e}^- \rightarrow \text{Be}$	-1.85	$\text{Cr}^{2+} + 2\text{e}^- \rightarrow \text{Cr}$	-0.91
$\text{Bi}^{3+} + 3\text{e}^- \rightarrow \text{Bi}$	+0.20	$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	+1.33
$\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^-$	+1.09	$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.74
$\text{BrO}^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Br}^- + 2\text{OH}^-$	+0.76	$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$	-0.41

$\text{Cs}^+ + \text{e}^- \rightarrow \text{Cs}$	-2.92	$\text{MnO}_4^- + 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{MnO}_2 + 4\text{OH}^-$	+0.60
$\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$	+0.52	$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71
$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.34	$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.23
$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	+0.16	$\text{NiOOH} + \text{H}_2\text{O} + \text{e}^- \rightarrow \text{Ni}(\text{OH})_2 + \text{OH}^-$	+0.49
$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	+2.87	$\text{NO}_3^- + 2\text{H}^+ + \text{e}^- \rightarrow \text{NO}_2 + \text{H}_2\text{O}$	-0.80
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.44	$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	+0.96
$\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}$	-0.04	$\text{NO}_3^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{NO}_2^- + 2\text{OH}^-$	+0.10
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	+0.77	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	+0.40
$[\text{Fe}(\text{CN})_6]^{3-} + \text{e}^- \rightarrow [\text{Fe}(\text{CN})_6]^{4-}$	+0.36	$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.23
$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0, by definition	$\text{O}_2 + \text{e}^- \rightarrow \text{O}_2^-$	-0.56
$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.83	$\text{O}_2 + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{HO}_2^- + \text{OH}^-$	-0.08
$2\text{HBrO} + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{Br}_2 + 2\text{H}_2\text{O}$	+1.60	$\text{O}_3 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{O}_2 + \text{H}_2\text{O}$	+2.07
$2\text{HClO} + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{Cl}_2 + 2\text{H}_2\text{O}$	+1.63	$\text{O}_3 + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{O}_2 + 2\text{OH}^-$	+1.24
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.78		
$\text{H}_4\text{XeO}_6 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{XeO}_3 + 3\text{H}_2\text{O}$	+3.0		
$\text{H}_4\text{XeO}_6 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{XeO}_3 + 3\text{H}_2\text{O}$	+3.0	$\text{O}_3 + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{O}_2 + 2\text{OH}^-$	+1.24
$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	+0.79	$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.13
$\text{Hg}_2\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Hg} + 2\text{Cl}^-$	+0.27	$\text{Pb}^{4+} + 2\text{e}^- \rightarrow \text{Pb}^{2+}$	+1.67
$\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}$	+0.86	$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.36
$2\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	+0.92	$\text{Pt}^{2+} + 2\text{e}^- \rightarrow \text{Pt}$	+1.20
$\text{Hg}_2\text{SO}_4 + 2\text{e}^- \rightarrow 2\text{Hg} + \text{SO}_4^{2-}$	+0.62	$\text{Pu}^{4+} + \text{e}^- \rightarrow \text{Pu}^{3+}$	+0.97
$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	+0.54	$\text{Ra}^{2+} + 2\text{e}^- \rightarrow \text{Ra}$	-2.92
$\text{I}_3^- + 2\text{e}^- \rightarrow 3\text{I}^-$	+0.53	$\text{Rb}^+ + \text{e}^- \rightarrow \text{Rb}$	-2.93
$\text{In}^+ + \text{e}^- \rightarrow \text{In}$	-0.14	$\text{S} + 2\text{e}^- \rightarrow \text{S}^{2-}$	-0.48
$\text{In}^{2+} + \text{e}^- \rightarrow \text{In}^+$	-0.40	$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}$	+2.05
$\text{In}^{3+} + 2\text{e}^- \rightarrow \text{In}^+$	-0.44	$\text{SC}^{3+} + 3\text{e}^- \rightarrow \text{Sc}$	-2.09
$\text{In}^{3+} + 3\text{e}^- \rightarrow \text{In}$	-0.34	$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.14
$\text{In}^{3+} + \text{e}^- \rightarrow \text{In}^{2+}$	-0.49	$\text{Sn}^{4+} + 2\text{e}^- \rightarrow \text{Sn}^{2+}$	+0.15
$\text{K}^+ + \text{e}^- \rightarrow \text{K}$	-2.93	$\text{Sr}^{2+} + 2\text{e}^- \rightarrow \text{Sr}$	-2.89
$\text{La}^{3+} + 3\text{e}^- \rightarrow \text{La}$	-2.52	$\text{Ti}^{2+} + 2\text{e}^- \rightarrow \text{Ti}$	-1.63
$\text{Li} + \text{e}^- \rightarrow \text{Li}$	-3.05	$\text{Ti}^{3+} + \text{e}^- \rightarrow \text{Ti}^{2+}$	-0.37
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.36	$\text{Ti}^{4+} + \text{e}^- \rightarrow \text{Ti}^{3+}$	0.00
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.18	$\text{Tl}^+ + \text{e}^- \rightarrow \text{Tl}$	-0.34
$\text{Mn}^{3+} + \text{e}^- \rightarrow \text{Mn}^{2+}$	+1.51	$\text{U}^{3+} + 3\text{e}^- \rightarrow \text{U}$	-1.79
$\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}$	+1.23	$\text{U}^{4+} + \text{e}^- \rightarrow \text{U}^{3+}$	-0.61
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.51	$\text{V}^{2+} + 2\text{e}^- \rightarrow \text{V}$	-1.19
$\text{MnO}_4^- + \text{e}^- \rightarrow \text{MnO}_4^{2-}$	+0.56	$\text{V}^{3+} + \text{e}^- \rightarrow \text{V}^{2+}$	-0.26
		$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76