Oxidation-Reduction Reactions

LEO SAYS GER

Oxidation and Reduction (Redox)

- Electrons are transferred
- Spontaneous redox rxns can transfer energy
    - Electrons (electricity)
    - Heat
- Non-spontaneous redox rxns can be made to happen with electricity

Oxidation Reduction Reactions (Redox)

$$\overset{0}{2\,Na} + \overset{0}{Cl_2} \rightarrow 2\,\overset{+1}{Na}\overset{-1}{Cl}$$

Each sodium atom loses one electron:

$$\overset{0}{Na} \rightarrow \overset{+1}{Na} + e^-$$

Each chlorine atom gains one electron:

$$\overset{0}{Cl} + e^- \rightarrow \overset{-1}{Cl}$$

LEO says GER :

Lose **E**lectrons = **O**xidation

$$\overset{0}{Na} \rightarrow \overset{+1}{Na} + e^{-}$$

Sodium is oxidized

Gain **E**lectrons = **R**eduction

$$\overset{0}{Cl} + e^{-} \rightarrow \overset{-1}{Cl}$$

Chlorine is reduced

Not All Reactions are Redox Reactions

Reactions in which there has been no change in oxidation number are not redox rxns.

Examples:

$$\overset{+1}{Ag}\,\overset{+5}{N}\,\overset{-2}{O}_3(aq) + \overset{+1}{Na}\,\overset{-1}{Cl}(aq) \rightarrow \overset{+1}{Ag}\,\overset{-1}{Cl}(s) + \overset{+1}{Na}\,\overset{+5}{N}\,\overset{-2}{O}_3(aq)$$

$$2\,\overset{+1}{Na}\,\overset{-2}{O}\,\overset{+1}{H}(aq) + \overset{+1}{H}_2\,\overset{+6}{S}\,\overset{-2}{O}_4(aq) \rightarrow +\,\overset{+1}{Na}_2\,\overset{+6}{S}\,\overset{-2}{O}_4(aq) + \overset{+1}{H}_2\,\overset{-2}{O}(l)$$

Rules for Assigning Oxidation Numbers
Rules 1 & 2

1. The oxidation number of any uncombined element is zero
2. The oxidation number of a monatomic ion equals its charge

$$2\overset{0}{Na} + \overset{0}{Cl_2} \rightarrow 2\overset{+1}{Na}\overset{-1}{Cl}$$

Rules for Assigning Oxidation Numbers
Rules 3 & 4

3. The oxidation number of oxygen in compounds is -2

4. The oxidation number of hydrogen in compounds is +1

$$\overset{+1}{H_2}\overset{-2}{O}$$

Rules for Assigning Oxidation Number
Rule 5

5. The sum of the oxidation numbers in the formula of a compound is 0

$$\overset{+1}{H}_2\overset{-2}{O}$$

2(+1) + (-2) = 0
H O

$$\overset{+2}{Ca}(\overset{-2}{O}\overset{+1}{H})_2$$

(+2) + 2(-2) + 2(+1) = 0
Ca O H

Rules for Assigning Oxidation Numbers
Rule 6

6. The sum of the oxidation numbers in the formula of a polyatomic ion is equal to its charge

$$\overset{?}{N}\overset{-2}{O}_3^{\,-}$$

$X + 3(-2) = -1$
N O

$\therefore X = +5$

$$\overset{?}{S}\overset{-2}{O}_4^{\,2-}$$

$X + 4(-2) = -2$
S O

$\therefore X = +6$

Reducing Agents and Oxidizing Agents

- The substance reduced is the oxidizing agent
- The substance oxidized is the reducing agent

$$\overset{0}{Na} \rightarrow \overset{+1}{Na} + e^{-}$$

Sodium is oxidized – it is the reducing agent

$$\overset{0}{Cl} + e^{-} \rightarrow \overset{-1}{Cl}$$

Chlorine is reduced – it is the oxidizing agent

Trends in Oxidation and Reduction

Active metals:

- Lose electrons easily
- Are easily oxidized
- Are strong reducing agents

Active nonmetals:

- Gain electrons easily
- Are easily reduced
- Are strong oxidizing agents

Electrochemical Terminology

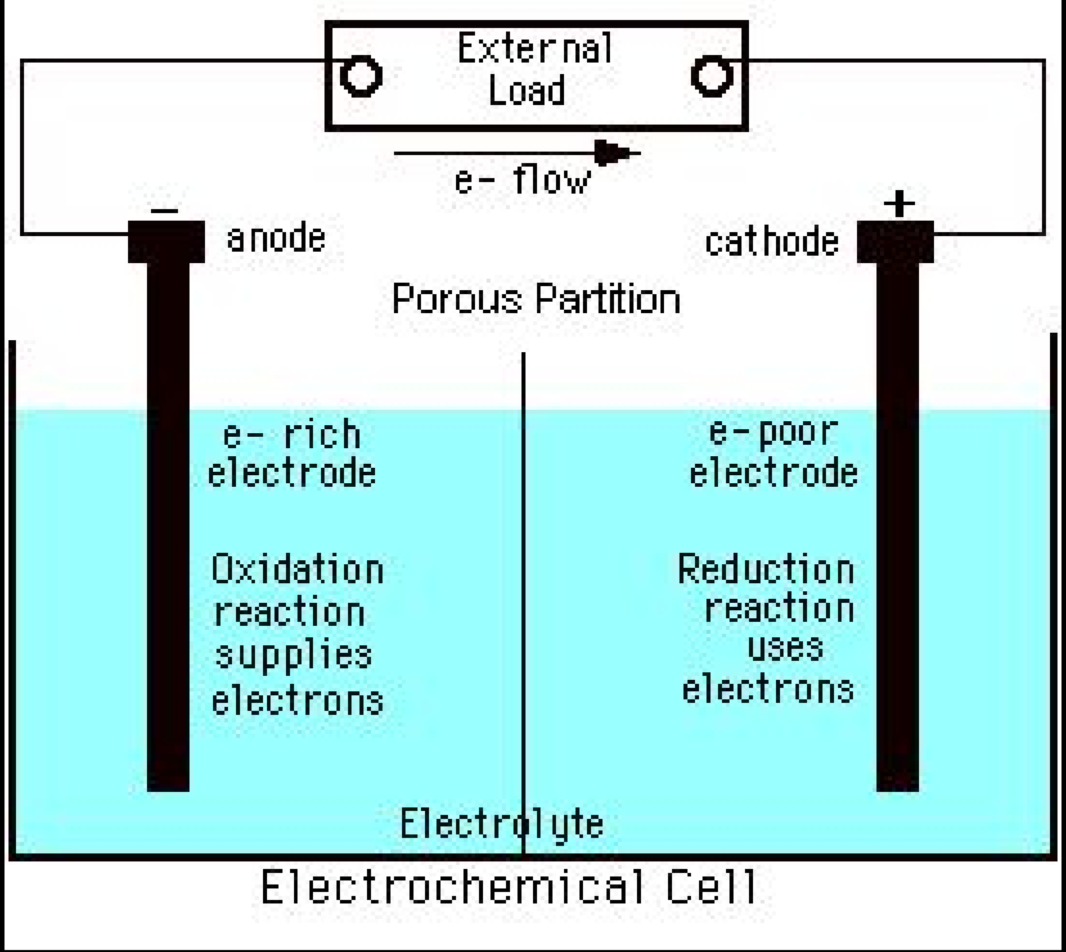
Electrode: A conductor used to establish contact with a nonmetallic part of a circuit, such as an electrolyte

Half-cell: a metal electrode in contact with a solution of its own ions

Anode: The electrode where oxidation takes place

Cathode: The electrode where reduction takes place

Voltaic Cells

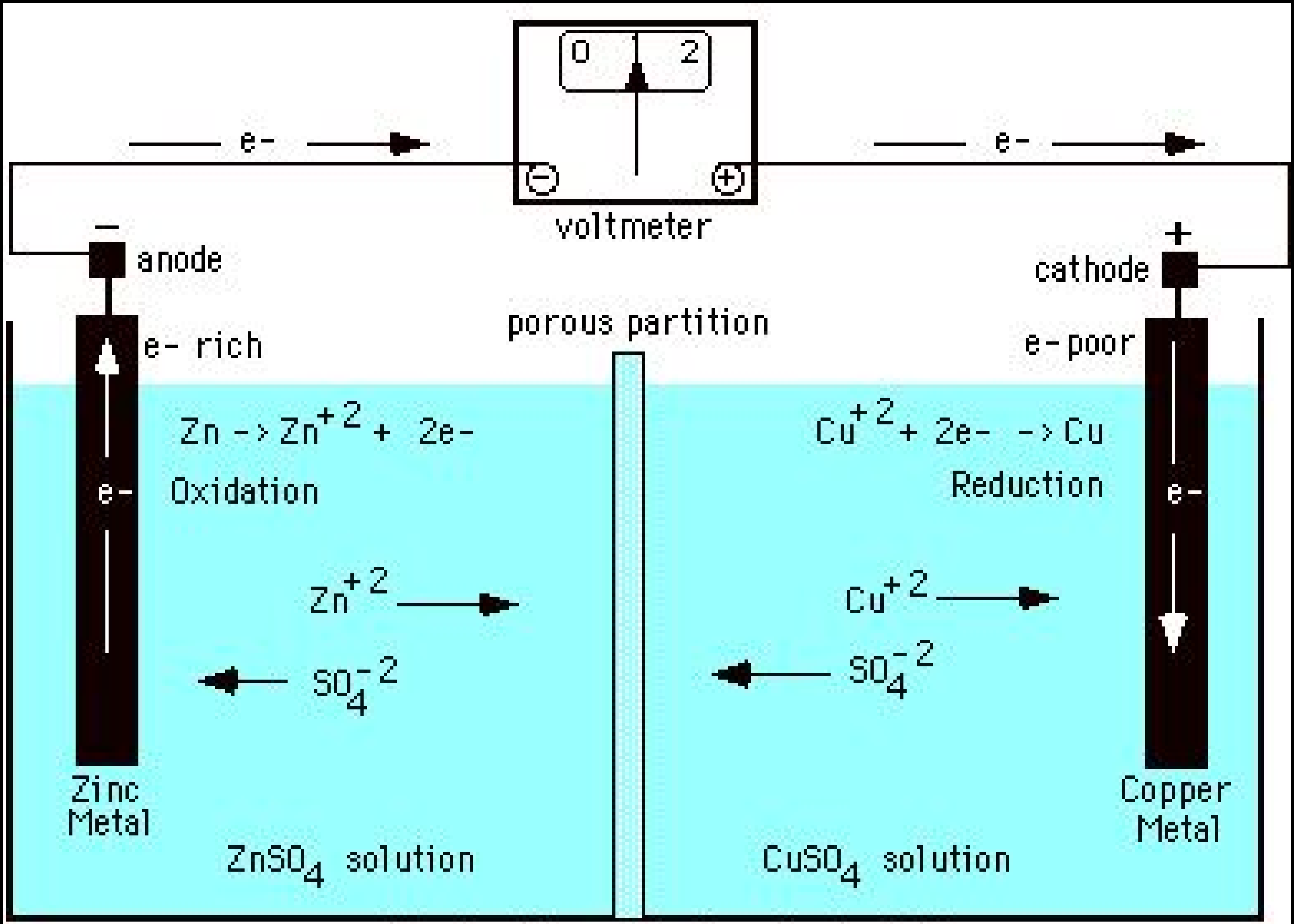


Anode: negative

Cathode: positive

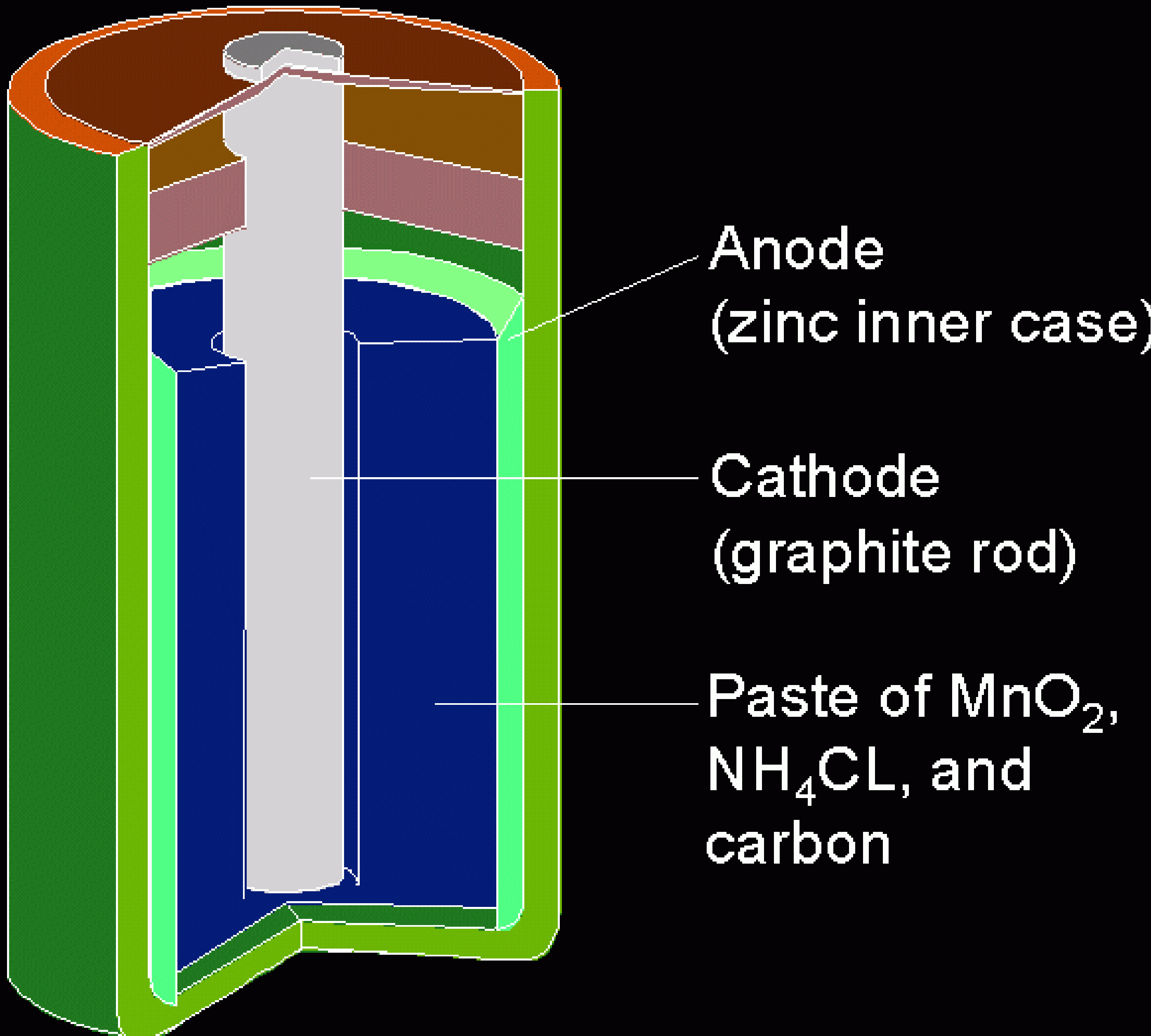
Voltaic cells: Electrochemical cells in which a spontaneous redox reaction can be harnessed to produce an electric current.

Zinc - Copper Battery

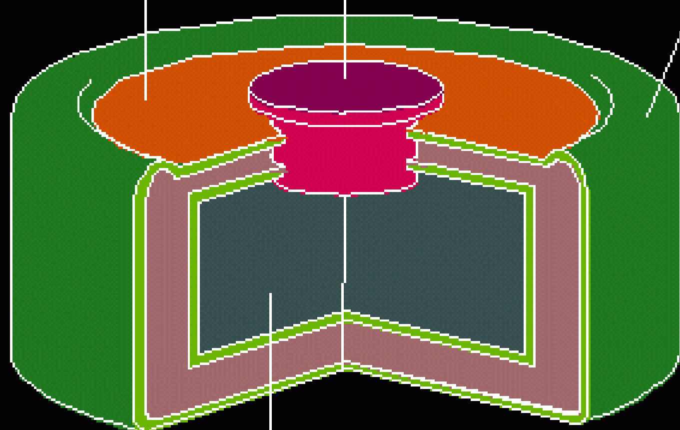


Zn-Cu Electrochemical Cell

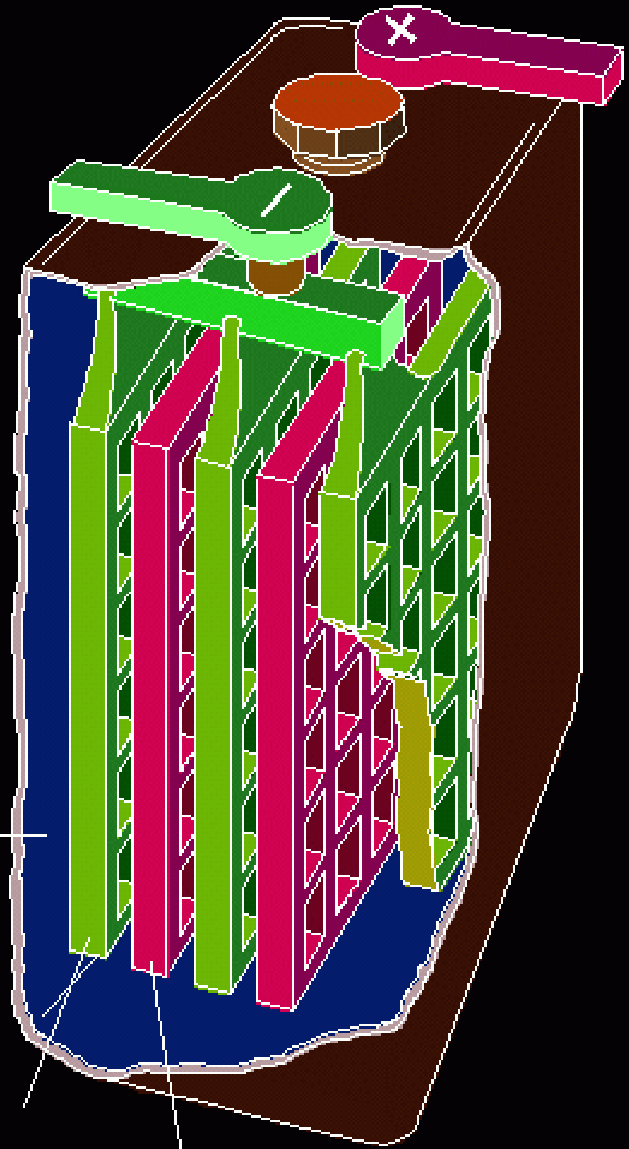
Zinc-Carbon Dry Cell Battery



Mercury Dry Cell Battery

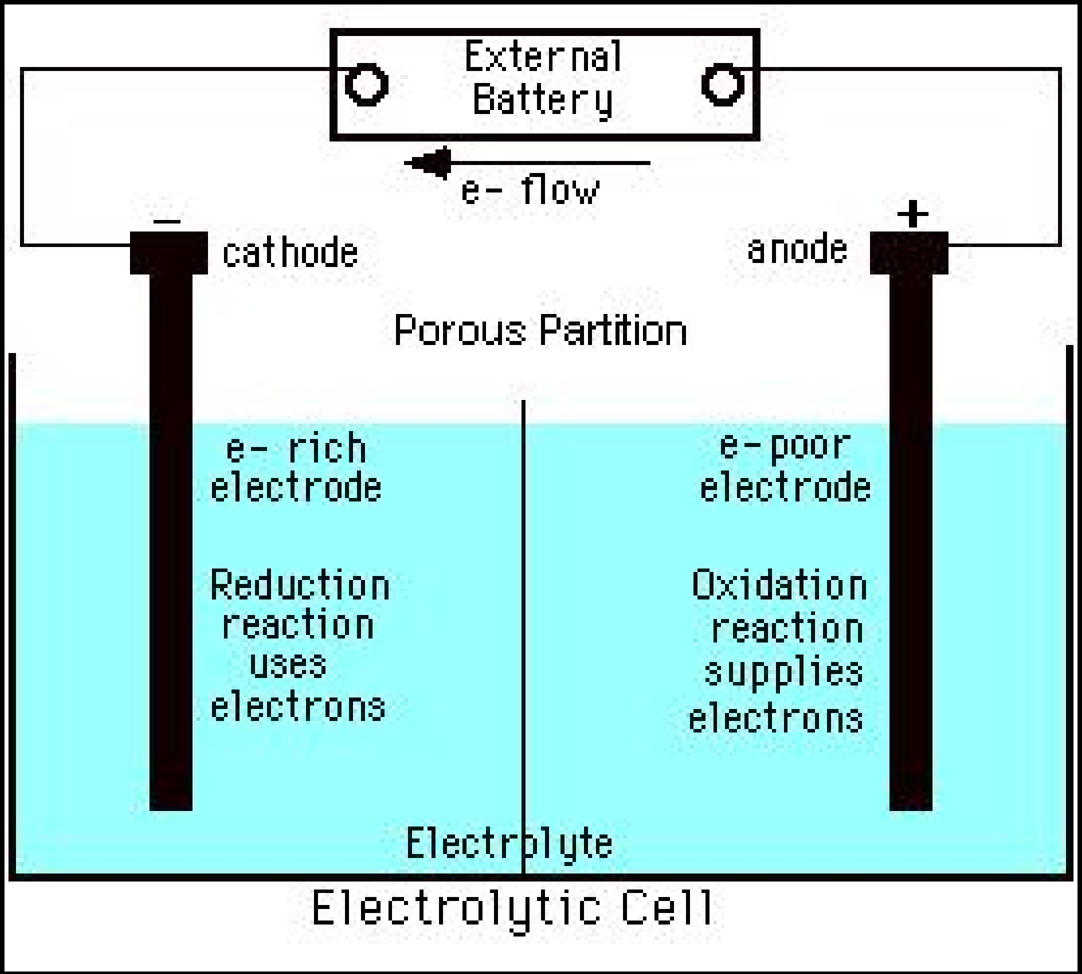


Lead Storage Automotive Battery



Electrolytic Cells

Cathode: negative



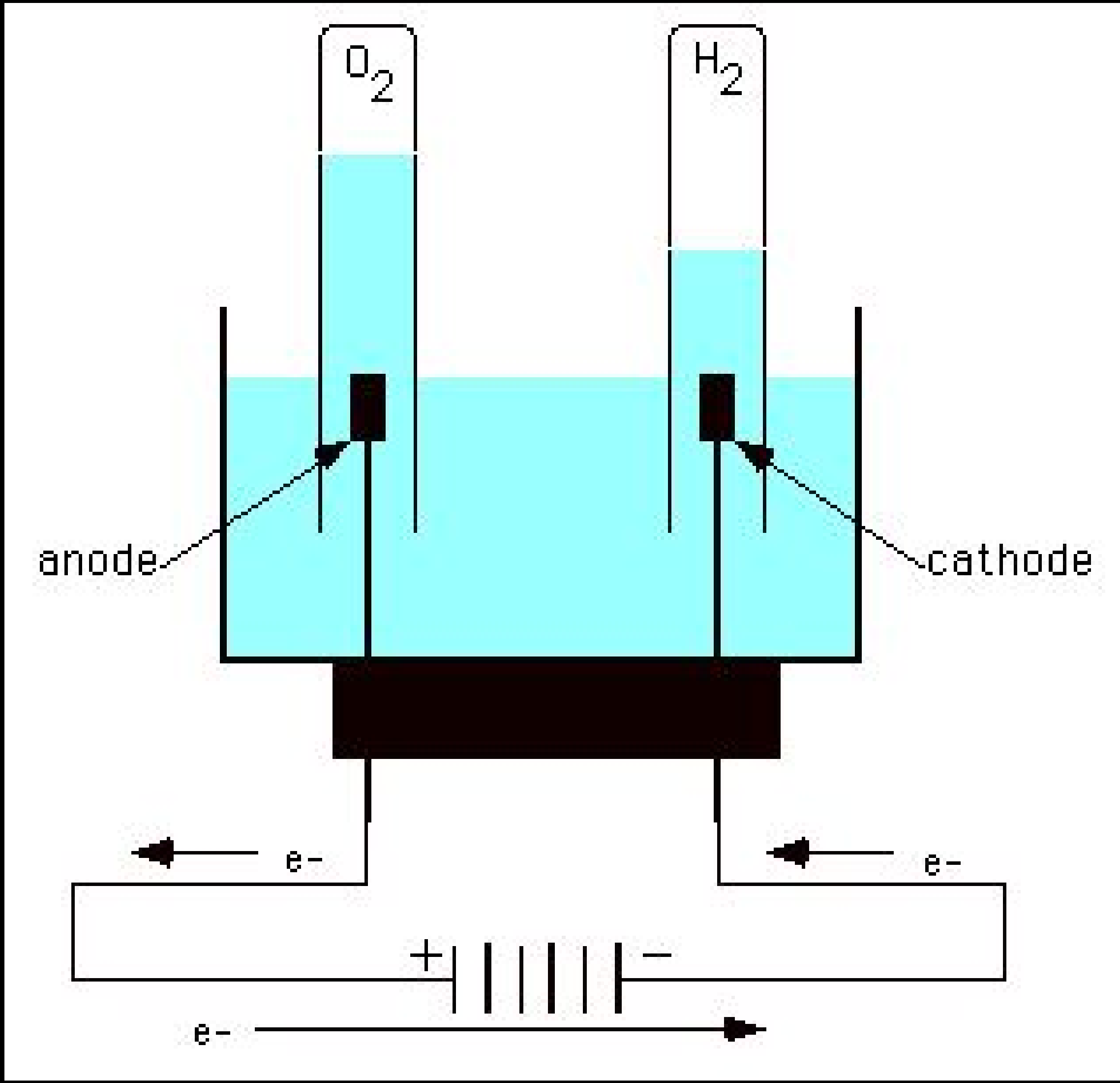
Anode: positive

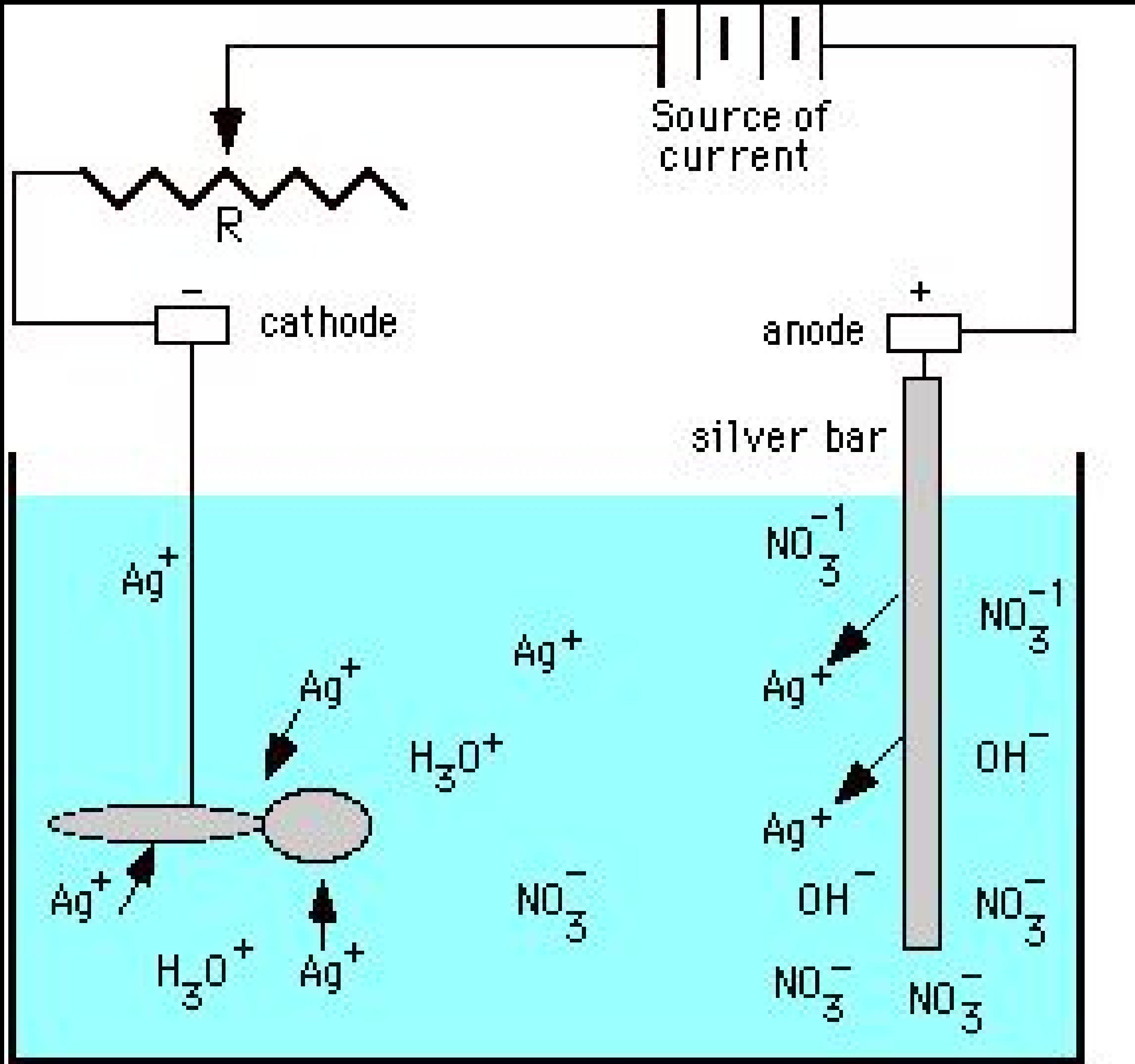
Electrolytic cells: Electrochemical cell in which an electric current is used to drive a non-spontaneous process

Electrolysis of H_2O

$2H_2O \rightarrow 2H_2 + O_2$

Electrolyte is usually dilute sulfuric acid





Electroplating of Silver

Cathode is the object to be plated

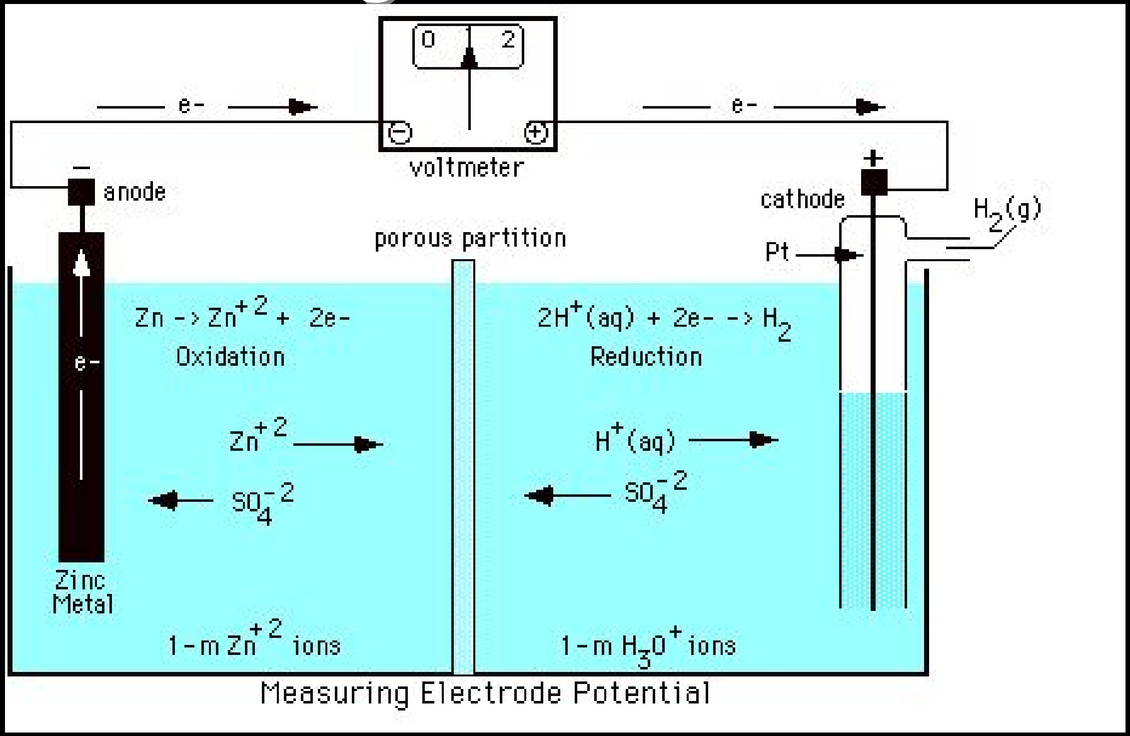
Anode is a piece of the plating metal

Solution contains ions of the plating metal

Table 17.1 Standard Reduction Potentials at 25°C (298 K) for Many Common Half-reactions

Half-reaction	ξ° (V)	Half-reaction	ξ° (V)
$F_2 + 2e^- \rightarrow 2F^-$	2.87	$O_2 + 2H_2O + 4e^- \rightarrow 4HO^-$	0.40
$Ag^{2+} + e^- \rightarrow Ag^+$	1.99	$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.82	$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	0.34
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78	$AgCl + e^- \rightarrow Ag + Cl^-$	0.22
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.70	$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2SO_3 + H_2O$	0.20
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$	1.69	$Cu^{2+} + e^- \rightarrow Cu^+$	0.16
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$	1.68	$2H^+ + 2e^- \rightarrow H_2$	0.00
$2e^- + 2H^+ + IO_4^- \rightarrow IO_3^- + H_2O$	1.60	$Fe^{3+} + 3e^- \rightarrow Fe$	−0.036
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51	$Pb^{2+} + 2e^- \rightarrow Pb$	−0.13
$Au^{3+} + 3e^- \rightarrow Au$	1.50	$Sn^{2+} + 2e^- \rightarrow Sn$	−0.14
$PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$	1.46	$Ni^{2+} + 2e^- \rightarrow Ni$	−0.23
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36	$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	−0.35
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33	$Cd^{2+} + 2e^- \rightarrow Cd$	−0.40
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23	$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.21	$Cr^{3+} + e^- \rightarrow Cr^{2+}$	−0.50
$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$	1.20	$Cr^{3+} + 3e^- \rightarrow Cr$	−0.73
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09	$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76
$VO_2 + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$	1.00	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	−0.83
$AuCl_4^- + 3e^- \rightarrow Au + 4Cl^-$	0.99	$Mn^{2+} + 2e^- \rightarrow Mn$	−1.18
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	0.96	$Al^{3+} + 3e^- \rightarrow Al$	−1.66
$ClO_2 + e^- \rightarrow ClO_2^-$	0.954	$H_2 + 2e^- \rightarrow 2H^-$	−2.23
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.91	$Mg^{2+} + 2e^- \rightarrow Mg$	−2.37
$Ag^+ + e^- \rightarrow Ag$	0.80	$La^{3+} + 3e^- \rightarrow La$	−2.37
$Hg_2^{2+} + 2e^- \rightarrow 2Hg^+$	0.80	$Na^+ + e^- \rightarrow Na$	−2.71
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77	$Ca^{2+} + 2e^- \rightarrow Ca$	−2.76
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68	$Ba^{2+} + 2e^- \rightarrow Ba$	−2.90
$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	0.56	$K^+ + e^- \rightarrow K$	−2.92
$I_2 + 2e^- \rightarrow 2I^-$	0.54	$Li^+ + e^- \rightarrow Li$	−3.05
$Cu^+ + e^- \rightarrow Cu$	0.52		

Measuring Electrode Potential



1. H_2 Electrode is assigned a potential of zero volts

2. Second electrode is placed in a solution of its own ions

3. Reduction potential of second electrode is measured.