

Chemistry

Electrochemistry

Grade 12

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Prepared by

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Student`s name :

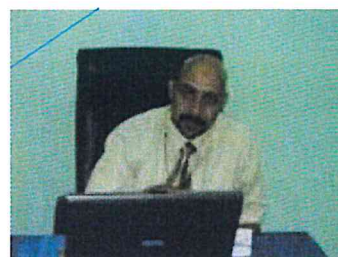
Class:

هذه المذكرة لاتعني عن الكتاب المدرسي وليست بهدف البيع او الربح

تمنيتاتي لجميع الطلبة والطالبات بالتفوق والنجاح

صدقة جارية علي روح كبير الكيميائيين

أ.سامي ابوالعلا رحمة الله عليه



Electrochemistry

Section 1 : Voltic Cells

Main idea :

- ❖ Chemical energy can be converted to electric energy and electric energy to chemical energy.
- ❖ In voltaic cells, oxidation takes place at the anode, yielding electrons that flow to the cathode, where reduction occurs.
- ❖ Voltaic cells have two half-cells, and both are required to produce energy.

items	Definition
Electrochemistry	is the study of the redox processes by which chemical energy is converted to electrical energy and vice versa
An electrochemical cell	is an apparatus that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction
A voltaic cell	is a type of electrochemical cell that converts chemical energy to electrical energy by a spontaneous redox reaction.
A salt bridge	is a pathway to allow the passage of ions from one side to another
An electrode	is an electrically conductive material, usually a metallic strip or graphite, that conducts electrons into and out of the solution in the half cell.
anode	that electrode where oxidation takes place
cathode	That electrode where reduction takes place
reduction potential	Is the tendency of a substance to gain electrons
The standard hydrogen electrode	consists of a small sheet of platinum immersed in a hydrochloric acid (HCl) solution that has a hydrogen-ion concentration of 1M. Hydrogen gas (H ₂), at a pressure of 1 atm is bubbled in and the temperature is maintained at 25°C
an electric current	A flow of charged particles

➤ Redox in Electrochemistry

- ✓ Electrochemical processes are useful in industry and critically important for biological functioning.
- ✓ all redox reactions involve a transfer of electrons from the species that is oxidized to the species that is reduced.

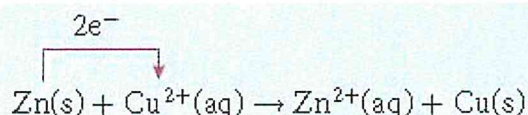
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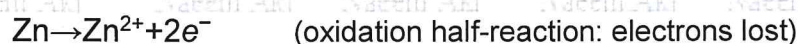
Example for the simple redox reaction

- in which zinc atoms are oxidized to form zinc (Zn^{2+}) ions.
- The two electrons donated from each zinc atom are accepted by a copper (Cu^{2+}) ion, which becomes an atom of copper metal.
- The following net ionic equation illustrates the electron transfer that occurs.



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Half-reactions



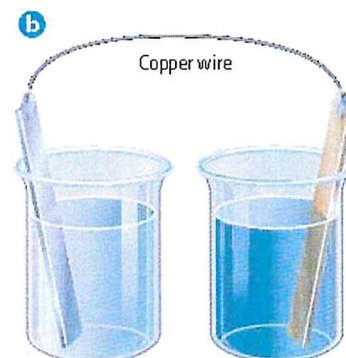
Activity 1

What do you think would happen if you separated the oxidation half-reaction from the reduction half-reaction?

in which a zinc strip is immersed in a solution of zinc sulfate and a copper strip is immersed in a solution of copper(II) sulfate.



- zinc metal is immersed in 1M zinc sulfate solution,
- and copper metal in 1M copper sulfate.
- Zinc will be oxidized
- while copper ions will be reduced



- wire joining the zinc and copper strips provides a pathway for the flow of electrons, but the pathway is not complete.
- Electron transfer is still not possible.

Two problems prevent a redox reaction from occurring:

(1) there is no way for zinc atoms to transfer electrons to copper(II) ions.

This problem can be solved by connecting the zinc and copper strips with **metal wire**. The wire can serve as a pathway for electrons to flow from the zinc strip to the copper strip.

(2) is that when the metal strips are placed in their solutions, oxidation begins at the zinc strip and reduction begins at the copper strip—but **these reactions cannot continue**. The reason is that as zinc is oxidized, positive zinc ions build up around the zinc electrode. As copper in the copper sulfate solution is reduced, negative sulfate ions build up around the copper electrode. The buildup of charges stops any further reaction. To solve this problem, **a salt bridge** must be added to the system.

➤ salt bridge

consists of a tube containing a conducting solution of a soluble salt, such as KCl, held in place by an agar gel or other permeable plug. Ions can move through the plug, but the solutions in the two beakers cannot mix.

Functions of salt bridge :

- 1- it maintains the electric al neutrality of solutions in two half cells because of the flow of ions (**ionic equilibrium** between two half cell)
- 2- it **completes the cell circuit** (**completes the pathway**) by connecting the solutions of two- half cells, without allowing them to mix with each other
- 3- it **prevent mixing** of the electrode solutions

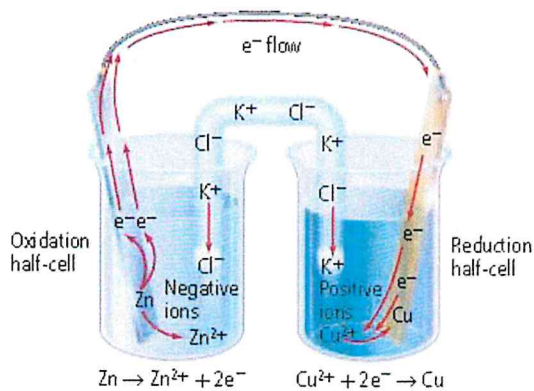
✓ When the connecting **metal wire** and **the salt bridge** are in place, the **spontaneous redox reaction begins**.

✓ Electrons flow through the wire from the oxidation half-reaction (**anode**) to the reduction half-reaction (**cathode**)

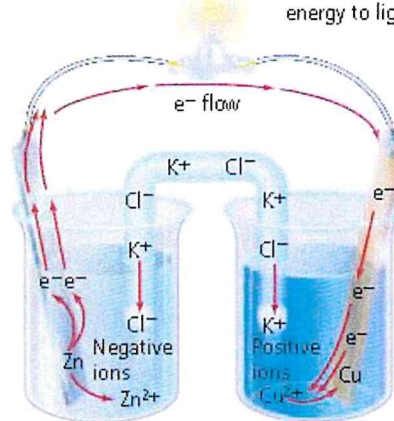
Notes

- ✓ while positive and negative ions move through the salt bridge.
- ✓ **Negative ions move through the salt bridge to the zinc side. Positive ions move through the bridge to the copper side**
- ✓ the flow of electrons through the wire and the flow of ions through the salt bridge make up the electric current.
- ✓ The energy of the flowing electrons can be used to light a bulb

The salt bridge and the wire provide an unbroken pathway for electrical charge to flow.



The redox reaction provides energy to light the bulb.



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Chemistry of Voltaic Cells

- ✓ An electrochemical cell consists of two parts called **half-cells**, in which the separate oxidation and reduction reactions take place.
- ✓ Each half-cell **contains** an **electrode** and a **solution containing ions**.
- ✓ the beaker with the zinc electrode is where the oxidation half of the redox reaction takes place.
- ✓ The beaker with the copper electrode is where the reduction half of the reaction takes place.
- ✓ The reaction that takes place in each halfcell is called a half-cell reaction.

Voltaic cells and energy

- ✓ electric potential energy is a measure of the amount of current that can be generated from a voltaic cell to do work.
- ✓ Electric charge can flow between two points only when a difference in electric potential energy exists between the two points.
- ✓ In an electrochemical cell, these two points are the two electrodes.
- ✓ Electrons generated at the anode, the site of oxidation, are thought to be pushed or driven toward the cathode by the electromotive force (EMF).
- ✓ This force is due to the difference in electric potential energy between the two electrodes and is referred to as **the cell potential**.

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✓ A volt is a unit used to measure cell potential.

✓ The electric potential difference of a voltaic cell is an indication of the energy that is available to move electrons from the anode to the cathode.

Example



- The roller coaster is momentarily stationary at the top of the track. Then, it plummets from its high position because of the difference in gravitational potential energy (PE) between the top and bottom of the track.
- The kinetic energy (KE) attained by the roller coaster is determined by the difference in height (potential energy) between the top and bottom parts of the track.
- **Similarly**, the energy of the electrons flowing from the **anode** to the **cathode** in a voltaic cell is determined by the difference in electric potential energy between the two electrodes.
- In redox terms, the voltage of a cell is determined by comparing the difference in **the tendency of the two electrode materials to accept electrons**. The greater the difference, the greater the potential energy difference between the two electrodes and the larger the voltage of the cell will be.

The roller coaster at the top of the track has high potential energy relative to track below because of the difference in height. **Similarly**, an electrochemical cell has potential energy to produce a current because there is a difference in the ability of the electrodes to move electrons from the anode to the cathode

Note

in the zinc-copper cell, under standard conditions, copper(II) ions **at the cathode accept electrons more readily than the zinc ions at the anode**. Thus, the redox reaction occurs spontaneously only when the electrons flow from the zinc to the copper.

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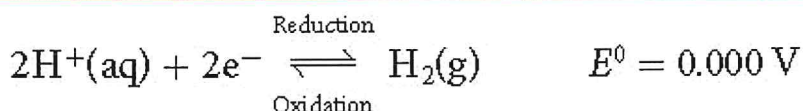
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➤ Calculating Electrochemical Cell Potentials:

- ✓ The reduction potential of an electrode cannot be determined directly **because** the reduction half-reaction must be coupled with an oxidation half-reaction.
- ✓ When two half-reactions are coupled, the voltage generated corresponds to the difference in potential between the half-reactions.
- ✓ The electrical potential difference between two points is expressed in volts(V).

➤ The standard hydrogen electrode :

- ✓ used to measure the reduction potential of all electrodes
- ✓ the standard reduction **potential (E^0)**, of this standard hydrogen electrode is defined as **0.000 V**.
- ✓ This electrode can act as an oxidation half-reaction or a reduction half-reaction, depending on the half-cell to which it is connected.
- ✓ The two possible reactions at the hydrogen electrode are the following.

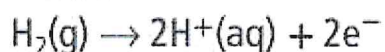


➤ Half-cell potentials

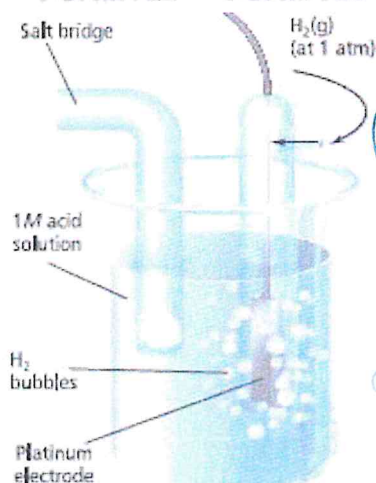
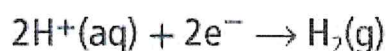
- ✓ The values in the table were obtained by measuring the potential when each half-cell was connected to a standard hydrogen half-cell.
- ✓ All of the half-reactions in **following Table** are written as reductions.
- ✓ However, in any voltaic cell, which always contains two half-reactions, the half-reaction with the **lower reduction potential** will proceed in the opposite direction and will be an oxidation reaction.
- ✓ In other words, the half-reaction that is more positive will proceed as a reduction and the half-reaction that is more negative will proceed as an oxidation.
- ✓ The electrode being measured must also be under **standard conditions**, that is, immersed in a **1M** solution of its ions at **25°C** and **1 atm**.
- ✓ The superscript zero in the notation E^0 is a shorthand way of indicating "**measured under standard conditions**."

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Oxidation:



Reduction:



A standard hydrogen electrode consists of a platinum electrode with hydrogen gas at 1 atm pressure bubbling into an acidic solution that is 1M in hydrogen ions. The reduction potential for this configuration is defined as 0.000 V.

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Table 1 Standard Reduction Potentials

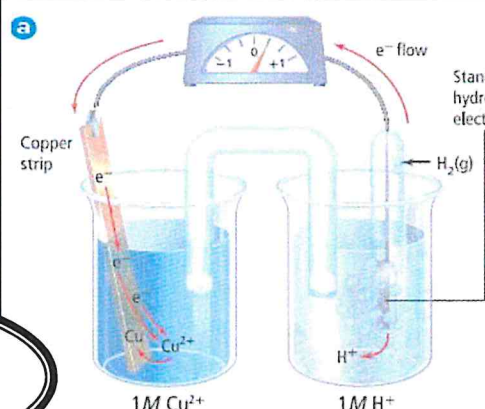
Half-Reaction	E^0 (V)	Half-Reaction	E^0 (V)
$\text{Li}^+ + \text{e}^- \rightleftharpoons \text{Li}$	-3.0401	$\text{Cu}^{2+} + \text{e}^- \rightleftharpoons \text{Cu}^+$	+0.153
$\text{Ca}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ca}$	-2.868	$\text{Cu}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cu}$	+0.3419
$\text{Na}^+ + \text{e}^- \rightleftharpoons \text{Na}$	-2.71	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightleftharpoons 4\text{OH}^-$	+0.401
$\text{Mg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Mg}$	-2.372	$\text{I}_2 + 2\text{e}^- \rightleftharpoons 2\text{I}^-$	+0.5355
$\text{Be}^{2+} + 2\text{e}^- \rightleftharpoons \text{Be}$	-1.847	$\text{Fe}^{3+} + \text{e}^- \rightleftharpoons \text{Fe}^{2+}$	+0.771
$\text{Al}^{3+} + 3\text{e}^- \rightleftharpoons \text{Al}$	-1.662	$\text{NO}_3^- + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{NO}_2 + \text{H}_2\text{O}$	+0.775
$\text{Mn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Mn}$	-1.185	$\text{Hg}_2^{2+} + 2\text{e}^- \rightleftharpoons 2\text{Hg}$	+0.7973
$\text{Cr}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cr}$	-0.913	$\text{Ag}^+ + \text{e}^- \rightleftharpoons \text{Ag}$	+0.7996
$2\text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{H}_2 + 2\text{OH}^-$	-0.8277	$\text{Hg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Hg}$	+0.851
$\text{Zn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Zn}$	-0.7618	$2\text{Hg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Hg}_2^{2+}$	+0.920
$\text{Cr}^{3+} + 3\text{e}^- \rightleftharpoons \text{Cr}$	-0.744	$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{NO} + 2\text{H}_2\text{O}$	+0.957
$\text{S} + 2\text{e}^- \rightleftharpoons \text{S}^{2-}$	-0.47627	$\text{Br}_2(\text{l}) + 2\text{e}^- \rightleftharpoons 2\text{Br}^-$	+1.066
$\text{Fe}^{2+} + 2\text{e}^- \rightleftharpoons \text{Fe}$	-0.447	$\text{Pt}^{2+} + 2\text{e}^- \rightleftharpoons \text{Pt}$	+1.18
$\text{Cd}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cd}$	-0.4030	$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$	+1.229
$\text{PbI}_2 + 2\text{e}^- \rightleftharpoons \text{Pb} + 2\text{I}^-$	-0.365	$\text{Cl}_2 + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-$	+1.35827
$\text{PbSO}_4 + 2\text{e}^- \rightleftharpoons \text{Pb} + \text{SO}_4^{2-}$	-0.3583	$\text{Au}^{3+} + 3\text{e}^- \rightleftharpoons \text{Au}$	+1.498
$\text{Co}^{2+} + 2\text{e}^- \rightleftharpoons \text{Co}$	-0.28	$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.507
$\text{Ni}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ni}$	-0.257	$\text{Au}^+ + \text{e}^- \rightleftharpoons \text{Au}$	+1.692
$\text{Sn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Sn}$	-0.1375	$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$	+1.776
$\text{Pb}^{2+} + 2\text{e}^- \rightleftharpoons \text{Pb}$	-0.1262	$\text{Co}^{3+} + \text{e}^- \rightleftharpoons \text{Co}^{2+}$	+1.92
$\text{Fe}^{3+} + 3\text{e}^- \rightleftharpoons \text{Fe}$	-0.037	$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightleftharpoons 2\text{SO}_4^{2-}$	+2.010
$2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2$	0.0000	$\text{F}_2 + 2\text{e}^- \rightleftharpoons 2\text{F}^-$	+2.866

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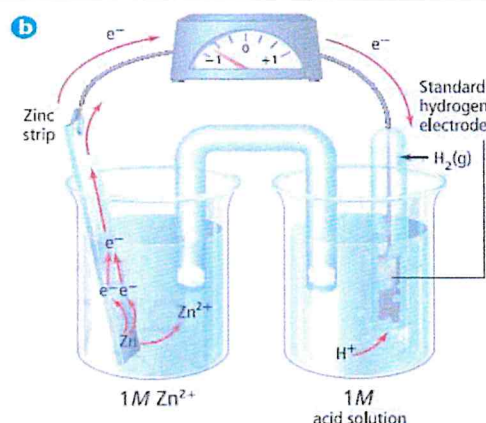
Determining electrochemical cell potentials

determine the standard reduction potential for the copper half-cell (E_{Cu}^0).



- When a $\text{Cu}|\text{Cu}^{2+}$ electrode is connected to the hydrogen electrode, electrons flow toward the copper strip and reduce Cu^{2+} ions to Cu atoms.
- electrons flow from the hydrogen electrode to the copper electrode, and copper ions are reduced to copper metal
- The voltage of this reaction is $+0.342 \text{ V}$
- The positive voltage indicates that Cu^{2+} ions at the copper electrode accept electrons more readily than do H^+ ions at the standard hydrogen electrode
- oxidation takes place at the hydrogen electrode and reduction takes place at the copper electrode.
- The oxidation and reduction half-cell reactions and the overall reaction are

determine the standard reduction potential for the zinc half-cell (E_{Zn}^0).



- When a $\text{Zn}|\text{Zn}^{2+}$ electrode is connected to the hydrogen electrode, electrons flow away from the zinc strip and zinc atoms are oxidized to Zn^{2+} ions.
- The voltage of this reaction is -0.762 V .
- electrons flow from the zinc electrode to the hydrogen electrode.
- The E^0 of the zinc half-cell, measured by a voltmeter, is -0.762 V .
- This means that the H^+ ions at the hydrogen electrode accept electrons more readily than do the zinc ions.
- the hydrogen ions have a higher reduction potential than the zinc ions.
- the reduction potential of the zinc electrode must have a negative value.

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Anode	Hydrogen electrode (H_2)	Zinc electrode (Zn)
Cathode	Copper electrode (Cu)	Hydrogen electrode (H_2)
Anode reaction	$\text{H}_2(\text{g}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{e}^-$	$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
Cathode reaction	$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$	$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$
overall redox reaction	$\text{H}_2(\text{g}) + \text{Cu}^{2+}(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + \text{Cu}(\text{s})$	$\text{Zn} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+} + \text{H}_2(\text{g})$
Cell notation	$\text{H}_2(\text{g}) 2\text{H}^+(\text{aq}) \text{Cu}^{2+}(\text{aq}) \text{Cu}(\text{s})$	$\text{Zn}(\text{s}) \text{Zn}^{2+}(\text{aq}) 2\text{H}^+(\text{aq}) \text{H}_2(\text{g})$

The final step in calculating electrochemical cell potential is to combine the copper and zinc half-cells as a voltaic cell.

Formula for Cell Potential

$$E_{\text{cell}}^0 = E_{\text{cathode}}^0 - E_{\text{anode}}^0$$

E_{cell}^0 represents the overall standard cell potential.

E_{cathode}^0 represents the standard half-cell potential for the reduction.

E_{anode}^0 represents the standard half-cell potential for the oxidation.

The standard potential of a cell is the standard potential of the reduction half-cell minus the standard potential of the oxidation half-cell.

Example 1

If galvanic cell consist of copper **Cu** and zinc electrode **Zn** answer the following ?

- Which is anode – cathode?
- Write anode reaction ?
- Write cathode reaction ?
- Write the over all reaction of the cell
- Write the cell notation ?
- Calculate the cell potential ?

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Solution :

✗ Anode is **Zn**

✗ cathode is **Cu**

✗ anode reaction: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$

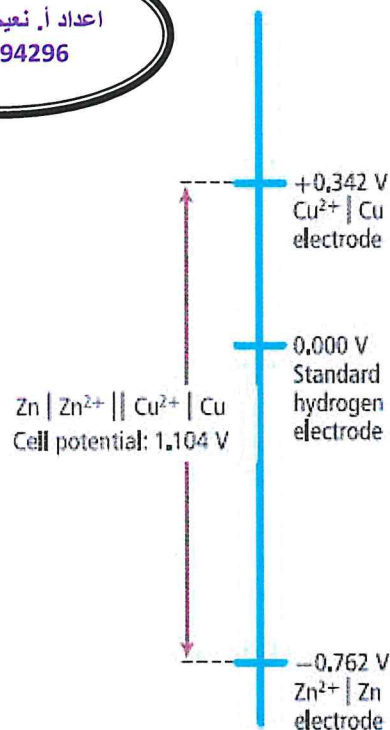
✗ cathode reaction $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$

✗ over all reaction of the cell $\text{Zn}_{(s)} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$

✗ the cell notation



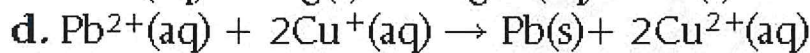
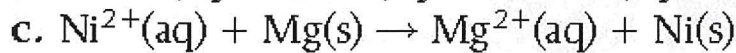
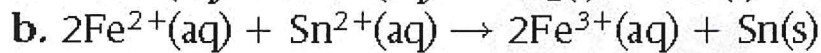
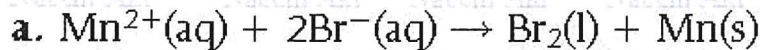
$$\begin{aligned} E_{\text{cell}}^0 &= E_{\text{Cu}^{2+}|\text{Cu}}^0 - E_{\text{Zn}^{2+}|\text{Zn}}^0 \\ &= +0.342 \text{ V} - (-0.762 \text{ V}) \\ &= +1.104 \text{ V} \end{aligned}$$



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Example 2

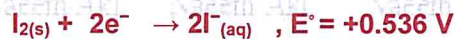
Determine whether each redox reaction is spontaneous or nonspontaneous.

**Notes :**

- that the negative sign in the formula automatically changes the sign of the oxidation half-reaction, so you do not reverse the sign of the standard reduction potentials listed in the Table when they are used for the oxidation half-reaction.
- the zinc half-cell with **the lower reduction potential is anode**
- and the copper half-cell with **the higher reduction potential is cathode**

Example 3

The following reduction half-reactions represent the half-cells of a voltaic cell.

**Determine**

- the overall cell reaction
- the standard cell potential.
- Describe the cell using cell notation.

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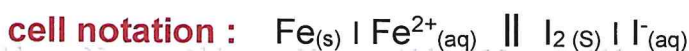
Solution :

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{reduction}} - E^{\circ}_{\text{oxidation}}$$

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{I}_2/\text{I}^{-}} - E^{\circ}_{\text{Fe}/\text{Fe}^{2+}}$$

$$E^{\circ}_{\text{cell}} = +0.536 \text{ V} - (-0.447 \text{ V})$$

$$E^{\circ}_{\text{cell}} = \mathbf{+0.983 \text{ V}}$$



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➤ Using Standard Reduction Potentials

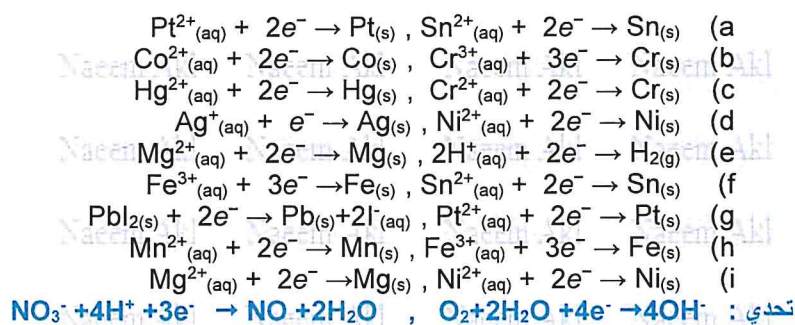
- 1- calculate the standard potential (**voltage**) of voltaic cells.
- 2- to determine which is **anode (low value)** and which is **cathode (higher value)**
- 3- To predict whether any proposed redox reaction will **occur spontaneously**:
 - ✎ If the calculated **potential is positive**, the reaction is **spontaneous**.
 - ✎ If the value is **negative**, the reaction is **not spontaneous**.

However, the reverse of a nonspontaneous reaction will occur because it will have a positive cell voltage, which means that the reverse reaction is spontaneous.

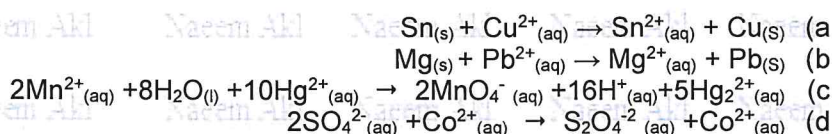
Applications :

- 1- For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction, and calculate the standard cell potential. Describe the reaction using cell notation.

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- 2- Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written.



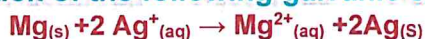
- 3- Challenge :write the equation and determine the cell voltage (E^0) for the following cell. Is the reaction spontaneous?



4- the standard potential for electrochemical cells in which each equation represents the overall cell reaction. Identify the reactions as spontaneous or nonspontaneous as written.

- $3\text{Cu}_{(s)} + 2\text{Al}^{3+}_{(aq)} \rightarrow 2\text{Cu}^{2+}_{(aq)} + 2\text{Al}_{(s)}$
- $\text{Hg}^{2+}_{(aq)} + 2\text{Cu}^{+}_{(aq)} \rightarrow 2\text{Cu}^{2+}_{(aq)} + \text{Hg}_{(l)}$
- $\text{Cd}_{(s)} + 2\text{NO}_3^{-}_{(aq)} + 4\text{H}^{+}_{(aq)} \rightarrow 2\text{NO}_{2(g)} + 2\text{H}_2\text{O}_{(l)} + \text{Cd}^{2+}_{(aq)}$

5- the over all reaction of the following galvanic cell is



- sketch the galvanic cell by represent the direction of electrons in wire ?
- Which is anode – cathode?
- Write anode reaction ?
- Write cathode reaction ?
- show direction of ion migration through the salt bridge
- Write the cell notation ?
- Calculate the cell potential ?

Selected Standard Reduction Potentials
at 25°C, 1 atm, and 1M Ion Concentration

Half-Reaction	E° (V)
$\text{Mg}^{2+} + 2\text{e}^{-} \rightarrow \text{Mg}$	-2.372
$\text{Al}^{3+} + 3\text{e}^{-} \rightarrow \text{Al}$	-1.662
$\text{Pb}^{2+} + 2\text{e}^{-} \rightarrow \text{Pb}$	-0.1262
$\text{Ag}^{+} + \text{e}^{-} \rightarrow \text{Ag}$	0.7996
$\text{Hg}^{2+} + 2\text{e}^{-} \rightarrow \text{Hg}$	0.851

6- Use the table below to answer Questions

1. Which metal ion is most easily reduced?

- ☐ Mg^{2+}
☐ Hg^{2+}
☐ Ag^{+}
☐ Al^{3+}

2. On the basis of the standard reduction potentials shown above, which standard cell notation correctly represents its voltaic cell?

- ☐ $\text{Ag} | \text{Ag}^{+} || \text{Al}^{3+} | \text{Al}$
☐ $\text{Mg} | \text{Mg}^{2+} || \text{H}^{+} | \text{H}_2$
☐ $\text{H}_2 | \text{H}^{+} || \text{Pb}^{2+} | \text{Pb}$
☐ $\text{Pb} | \text{Pb}^{2+} || \text{Al}^{3+} | \text{Al}$

3. A voltaic cell consists of a magnesium bar dipping into a 1M Mg^{2+} solution and a silver bar dipping into a 1M Ag^{+} solution.

What is the standard potential of this cell?

- ☐ 1.572 V
 ☐ 3.172 V
 ☐ 0.773 V
 ☐ 3.971 V

4. Assuming standard conditions, which cell will produce a potential of 2.513 V?

- ☐ $\text{Al} | \text{Al}^{3+} || \text{Hg}^{2+} | \text{Hg}$
☐ $\text{H}_2 | \text{H}^{+} || \text{Hg}^{2+} | \text{Hg}$
☐ $\text{Mg} | \text{Mg}^{2+} || \text{Al}^{3+} | \text{Al}$
☐ $\text{Pb} | \text{Pb}^{2+} || \text{Ag}^{+} | \text{Ag}$

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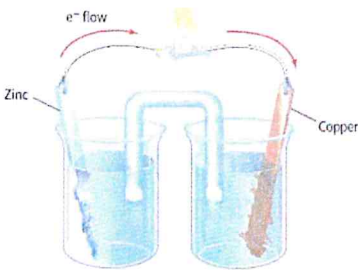
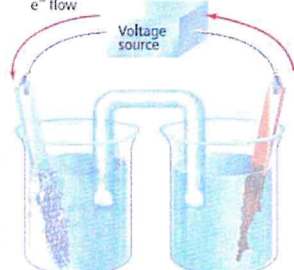
Section 3 : Electrolysis

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Main idea :

In electrolysis, a power source causes nonspontaneous reactions to occur in electrochemical cells.

items	Definition
electrolysis	The use of electrical energy to bring about a chemical reaction
electrolytic cell.	An electrochemical cell in which electrolysis occurs
Electroplating	Is electrochemical process that reduced the ions of metal from solution and deposited on the object to be plated

Voltic (galvanic) cell	Electrolytic (analytical) cell
 Voltaic cell	 Electrolytic cell
Convert chemical energy into electrical energy	Convert electric energy into chemical energy
spontaneous redox reaction.	Non spontaneous redox reaction.
Potential of cell is positive	Potential of cell is negative
Anode (-) , oxodaion	Anode (+) , oxidation
Cathode (+) , reduction	Cathode (-) , reduction
Not use an external voltage source .	using an external voltage source .
In this voltaic cell, the oxidation of zinc supplies the electrons to light the bulb and reduce copper ions. The spontaneous reaction continues until the zinc is used up.	When an outside voltage is applied, the flow of electrons is reversed and the nonspontaneous reaction occurs, which restores the conditions of the cell.

➤ Reversing Redox Reactions:

- ✓ When a battery generates electric current, electrons given up at the **anode** flow through an external circuit to the **cathode**, where they are used in a reduction reaction.
- ✓ One electrochemical cell is supplying power to a lightbulb by means of a spontaneous redox reaction
- ✓ Electrons flow spontaneously from the zinc side to the copper side, creating an electric current.
- ✓ The reaction continues until the zinc strip is used up, and then the reaction stops.
- ✓ However, **the cell can be regenerated** if current is applied in the reverse direction using an **external voltage source**.
- ✓ **A secondary battery** is one that can be recharged by passing a current through it in the opposite direction.
- ✓ when a secondary battery is recharged, it is acting as **an electrolytic cell**.
- ✓ **The voltage source** is required because the **reverse reaction** is nonspontaneous.
- ✓ If the voltage source is applied long enough, the cell will return to nearly its original strength

Applications of Electrolysis

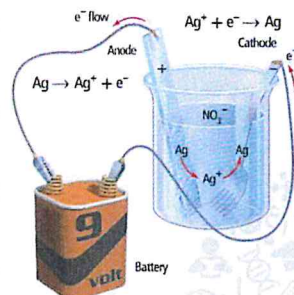
➤ Electroplating

Electroplating cell consist of :

- 1- **cathode of an electrolytic cell**: is **the object** to be plated
- 2- **The anode** is a **electroplating metal bar or sheet**
- 3- **Electrolyte** : **solution of electroplating metal**

Example :**electroplated with a metal such as silver**

- 1- **cathode of an electrolytic cell:** is **the object** to be plated (fork)
- 2- **The anode** is a **silver bar or sheet**
- 3- **Electrolyte :** **aqueous solution of silver nitrate AgNO_3**

Anode half- reaction:**Cathode half reaction :**

- **At the anode:** silver is oxidized to silver ions as electrons are removed by the power source.
- **At the cathode:** the silver ions are reduced to silver metal by electrons from the external power source and deposited on the object.
- The silver forms a thin coating over the object being plated.
- Current passing through the cell must be carefully controlled in order to get a smooth, even metal coating.
- Other metals are also used for electroplating like gold. Or nickel

Importance of electroplating :

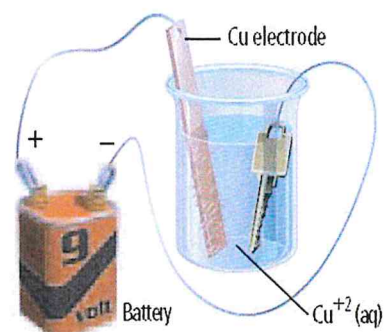
- Made metals **more corrosion-resistant** by being electroplated
For example : an automobile whose steel parts such as the bumper first electroplated with nickel and then with chromium.
- **Give metal high economic value**

⚡ Power is needed to oxidize silver at the anode and reduce silver at the cathode.

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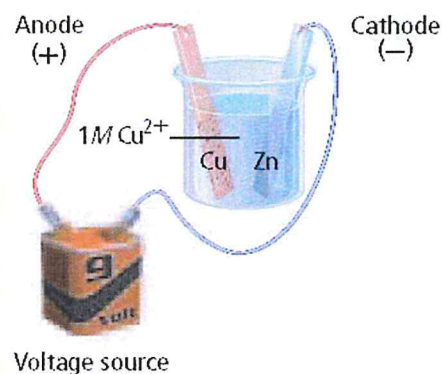
Applications :

- 1- In the below figure shows a key being electroplated with copper in an electrolytic cell. Where does oxidation occur? Explain your answer.



- 2- Answer the following questions based on following figure :

- Which electrode grows? Write the reaction that occurs at this electrode.
- Which electrode disappears? Write the reaction that occurs at this electrode.
- explain what happens to the copper ions in solution.



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- 3- By using the following tables answer the following ?

☞ which electrodes (Ag , Fe) is represent cathode in cell (1)

☞ which the following metals is the strongest reducing agent (Fe , Ag , Zn)

☞ which cell give high voltage ?.....

☞ which cell give low voltage ?.....

Electrode B	Electrode A	cell
Ag	Fe	1
Fe	Zn	2
Mg	Ag	3

E ⁰ (V)	Half- reaction cell
- 0.41	Fe ²⁺ + 2e ⁻ → Fe
- 0.76	Zn ²⁺ + 2e ⁻ → Zn
+ 0.80	Ag ⁺ + e ⁻ → Ag
- 2.37	Mg ²⁺ + 2e ⁻ → Mg

With my best wishes

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