CHEMISTRY

12 Advanced

Electrochemistry

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Sec. (1) Voltaic Cells

Electrochemistry: the study of oxidation and reduction processes in which chemical energy is converted into electrical energy and vice versa.

Electrochemical processes are useful in industry, and in biological functions.

Electrochemical Cells

Electrochemical cell: a device that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction.

Types of electrochemical cells

- **voltaic cell** → Convert chemical energy into electrical energy
- **electrolytic cell** → Convert electrical energy into chemical energy

Voltaic cell: a type of electrochemical cell that converts chemical energy into electrical energy through a spontaneous redox reaction.

An electrochemical cell consists of two parts called a half-cell.

Half-cell reaction: the reaction that takes place in each half of a cell.

The voltaic cell is named by the scientist (Alessandro Volta) and the corresponding shape resembles one of the first cells of Alessandro Volta.

The Volta cell consists of zinc and copper discs arranged in alternate layers, separated by pieces of cloth or cardboard dipped in an acidic solution.

The intensity of the current increases with the increase in the number of zinc and copper discs.

Electrode: a material that conducts electricity.

The electrode is a metallic strip or graphite stem that conducts electrons in and out of solution in the half-cell.

Anode: the electrode at which the oxidation reaction takes place.

Cathode: the electrode at which the reduction reaction takes place.

$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$

- half-reactions
  - oxidation (lose electrons): $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$
  - reduction (gain electrons): $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$
- $2e^-$ transfers from Zn to Cu^{2+}
  - the products: Zn^{2+} and Cu

But does this reaction occur when the oxidation half-reaction is separated from the reduction half-reaction?

zinc electrode in1M zinc sulfate solution

copper electrode in 1M copper sulfate solution

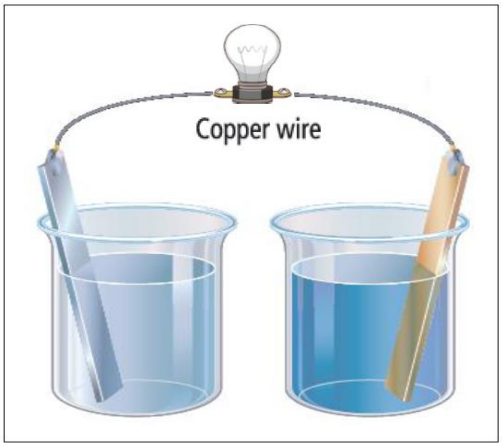
There is a problem? → There is no way to transfer electrons from Zn atoms to Cu^{2+} ions

Solve the problem → The zinc and copper electrodes are connected to a metal wire, where the wire acts as a path for the transmission of electrons.

What happens when the electrodes are connected to a metallic wire?

The lamp does not light

The metallic wire connecting the zinc and copper electrodes provides a path for the transfer of electrons, but the transfer of electrons does not occur.



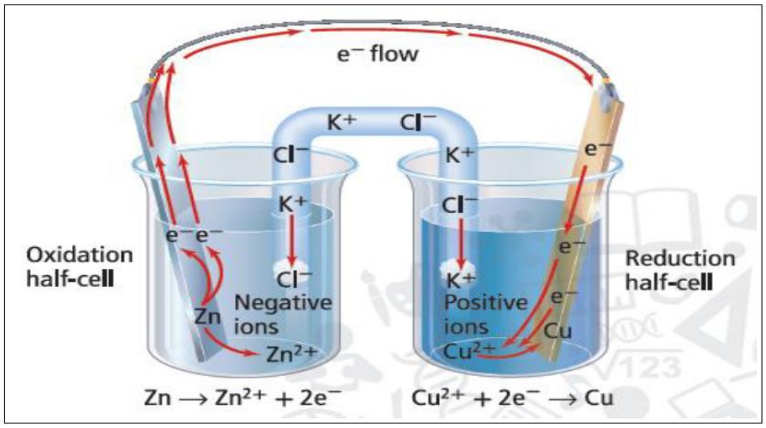
There is another problem? → **Zn^{2+} ions accumulate around the Zn electrode, Cu^{2+} ions are accumulate around the Cu electrode, and the charge accumulation prevents the reaction from continuing.**

Solve the problem → **Use a salt bridge.**

Salt Bridge: a path to maintain the neutrality of a solution by allowing the passage of ions from one side to the other.

What is a salt bridge made of?

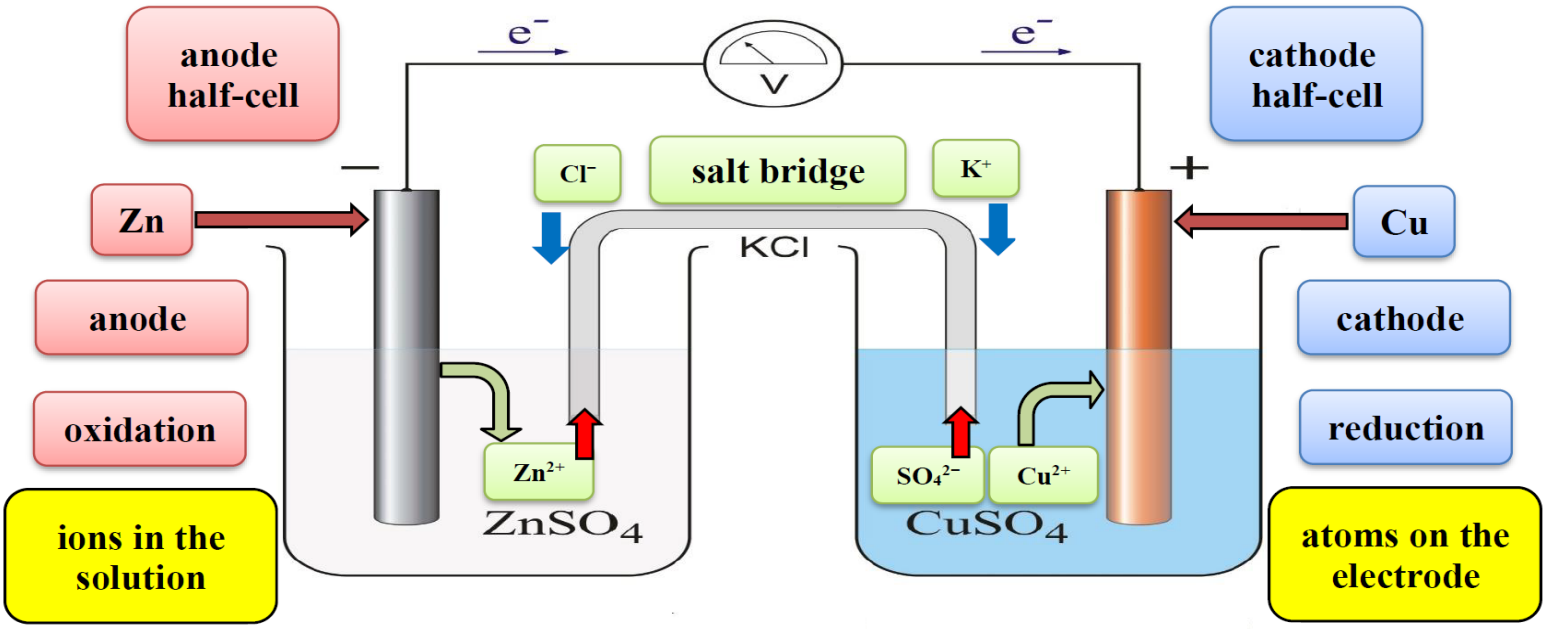
It consists of a tube containing a conductive salt solution such as KCl that holds the salt in place using an agar gel.



The agar gel in the brine allows the ions to move through it, while not allowing the two solutions to mix.

salt bridge functions

- **allows ions to pass from one side to the other**
- **restore ionic balance in the cell**
- **path to maintain solution neutrality**



if the metal wire and the salt bridge are in place

the oxidation and reduction process starts spontaneously

oxidation: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

Zn atoms turns to Zn^{2+} ions

electrons travel through the metal wire to the Cu electrode

reduction: $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

Cu^{2+} ions are reduced by electrons

$CuSO_4$ solution dissociates into Cu^{2+} and SO_4^{2-}

There is a problem?

The concentration of Zn^{2+} ions increases in half-cell of the anode, the concentration of SO_4^{2-} ions increases in half-cell of the cathode.

The role of the salt bridge comes

an exchange of ions occurs in half of the anode cell.

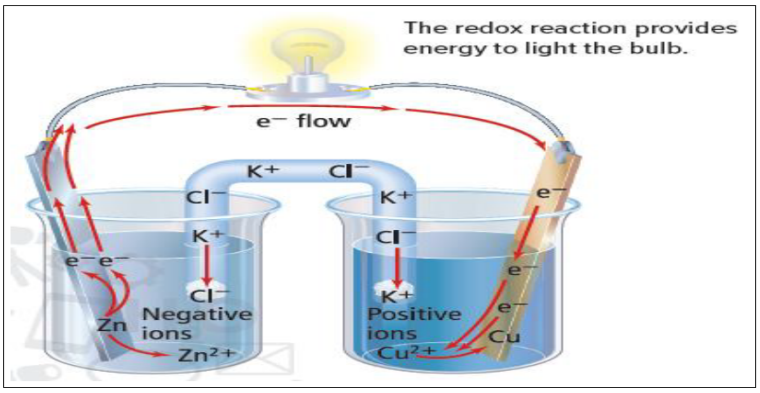
positive ions enter the salt bridge and negative ions leave it.

an exchange of ions occurs in half of the cathode cell.

negative ions enter the salt bridge and positive ions leave it.

What results from the flow of electrons through the wire, and the flow of ions through the salt bridge?

Electric current.



How do the electrode masses and ion concentrations change in the cell?

- **anode electrode mass decreases (Zn)**
- **cathode electrode mass increases (Cu)**
- **anode solution ion concentration increases (Zn^{2+})**
- **cathode solution ion concentration decreases (Cu^{2+})**

The movement of electrons and ions in the cell

- **Electrons move through the metal wire from the anode to the cathode.**
- **Positive ions move across the salt bridge (from the anode half-cell towards the the cathode half-cell).**
- **Negative ions move across the salt bridge (from the cathode half-cell towards the anode half-cell).**

Oxidation half-reaction	$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$
Reduction half-reaction	$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$
Overall reaction	$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$

	anode	cathode
Cell notation	reactant | product	reactant | product
Cell notation	Zn | Zn^{2+}	Cu^{2+} | Cu

A voltaic cell expressed by the reaction: $Pb^{2+} + Mg \rightarrow Pb + Mg^{2+}$,

answer the questions (1-3)

1) **What is the equation for the reaction of the anode?**

   a. $Mg^{2+} + 2e^- \rightarrow Mg$

   b. $Mg \rightarrow Mg^{2+} + 2e^-$

   c. $Pb^{2+} + 2e^- \rightarrow Pb$

   d. $Pb \rightarrow Pb^{2+} + 2e^-$

2) **What is the equation for the reaction of the cathode?**

   a. $Mg^{2+} + 2e^- \rightarrow Mg$

   b. $Mg \rightarrow Mg^{2+} + 2e^-$

   c. $Pb^{2+} + 2e^- \rightarrow Pb$

   d. $Pb \rightarrow Pb^{2+} + 2e^-$

3) **What is the cell notation?**

   a. $Pb^{2+} \mid Pb \parallel Mg \mid Mg^{2+}$

   b. $Mg^{2+} \mid Mg \parallel Pb \mid Pb^{2+}$

   c. $Pb \mid Pb^{2+} \parallel Mg^{2+} \mid Mg$

   d. $Mg \mid Mg^{2+} \parallel Pb^{2+} \mid Pb$

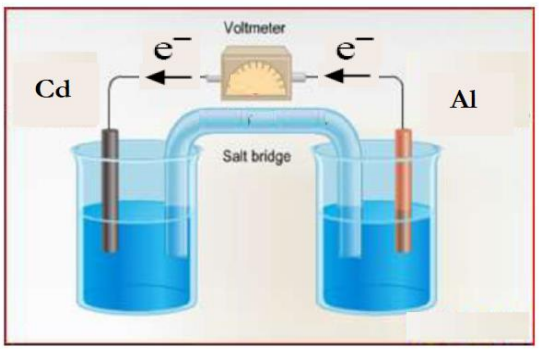
4) **Depending on the corresponding cell, what reaction occurs at the anode?**

   a. $Al \rightarrow Al^{3+} + 3e^-$

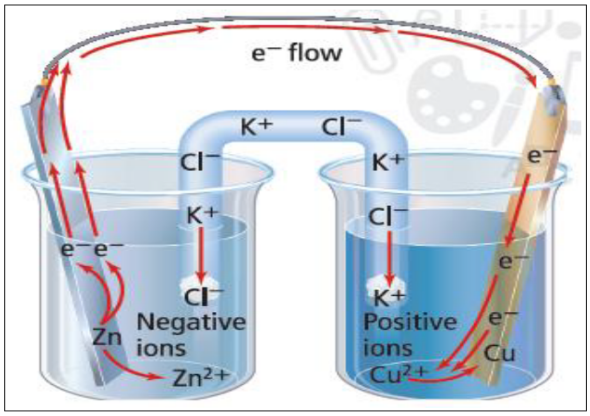
   b. $Al^{3+} + 3e^- \rightarrow Al$

   c. $Cd \rightarrow Cd^{2+} + 2e^-$

   d. $Cd^{2+} + 2e^- \rightarrow Cd$



Using the corresponding figure, answer the questions (5-7).



5) What is the anode reaction?

a. $Cu \rightarrow Cu^{2+} + 2e^-$ b. $Cu^{2+} + 2e^- \rightarrow Cu$

c. $Zn \rightarrow Zn^{2+} + 2e^-$ d. $Zn^{2+} + 2e^- \rightarrow Zn$

6) What happens to the cathode electrode mass?

a. increase b. decrease

c. do not change d. can not be determined

7) What is the cell notation?

a. $Cu \mid Cu^{2+} \parallel Zn^{2+} \mid Zn$ b. $Cu^{2+} \mid Cu \parallel Zn \mid Zn^{2+}$

c. $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$ d. $Zn^{2+} \mid Zn \parallel Cu \mid Cu^{2+}$

8) In the voltaic cell represented by: $Al \mid Al^{3+} \parallel Cu^{2+} \mid Cu$, what is oxidized and what is reduced when the cell releases current?

a. oxidation: Al^{3+}, reduction: Cu b. oxidation: Al, reduction: Cu^{2+}

c. oxidation: Cu^{2+}, reduction: Al d. oxidation: Cu, reduction: Al^{3+}

9) What is the equation for the reaction at the cathode of the voltaic cell in which the reaction takes place? $Ba + Sn^{2+} \rightarrow Ba^{2+} + Sn$

a. $Ba \rightarrow Ba^{2+} + 2e^-$ b. $Ba^{2+} + 2e^- \rightarrow Ba$

c. $Sn \rightarrow Sn^{2+} + 2e^-$ d. $Sn^{2+} + 2e^- \rightarrow Sn$

10) A voltaic cell using chromium and iron (II) is constructed as follows, what phrase describes this system? $2Cr_{(s)} + 3Fe^{2+}_{(aq)} \rightarrow 2Cr^{3+}_{(aq)} + 3Fe_{(s)}$

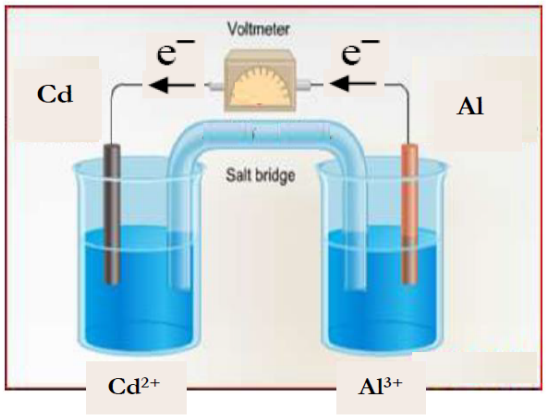
a. electrons flow from the iron electrode to the chromium electrode

b. heat energy is released

c. negative ions move across the salt bridge from the chromium half-cell to the iron half-cell

d. negative ions move across the salt bridge from the iron half-cell to the chromium half-cell

Using the corresponding figure, answer questions (11-14)



11) What is the reaction of the anode?

a. $Cd^{2+} + 2e^- \rightarrow Cd$

b. $Cd \rightarrow Cd^{2+} + 2e^-$

c. $Al^{3+} + 3e^- \rightarrow Al$

d. $Al \rightarrow Al^{3+} + 3e^-$

12) What is the reaction of the cathode?

a. $Cd^{2+} + 2e^- \rightarrow Cd$

b. $Cd \rightarrow Cd^{2+} + 2e^-$

c. $Al^{3+} + 3e^- \rightarrow Al$

d. $Al \rightarrow Al^{3+} + 3e^-$

13) What is the overall reaction of the cell?

a. $2Al^{3+} + 3Cd \rightarrow 2Al + 3Cd^{2+}$

b. $2Al + 3Cd^{2+} \rightarrow 2Al^{3+} + 3Cd$

c. $2Al + 3Cd \rightarrow 2Al^{3+} + 3Cd^{2+}$

d. $2Al^{3+} + 3Cd^{2+} \rightarrow 2Al + 3Cd$

14) What is the cell notation?

a. $Al^{3+} \mid Al \parallel Cd \mid Cd^{2+}$

b. $Cd \mid Cd^{2+} \parallel Al^{3+} \mid Al$

c. $Al \mid Al^{3+} \parallel Cd^{2+} \mid Cd$

d. $Cd^{2+} \mid Cd \parallel Al \mid Al^{3+}$

15) Which of the following options occurs in the voltaic cell that has the notation?

$Al \mid Al^{3+} \parallel Fe^{2+} \mid Fe$

a. reduction: Fe^{2+}

b. reduction: Al^{3+}

c. reduction: Al

d. oxidation: Fe

11	d	12	a	13	b	141	c	15	a

Voltaic cells and energy

The potential energy of the electrode depends on its position or configuration.

When can an electric charge move between two points? → **If there is a difference in electrical energy between them.**

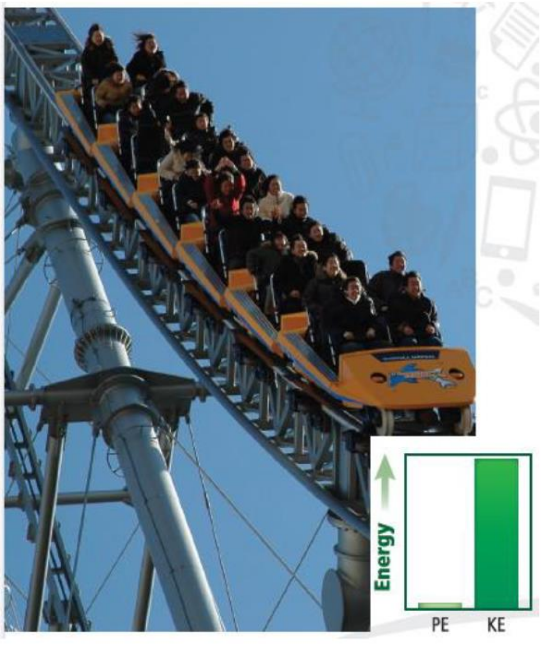
What are the two points between which charge moves in an electrochemical cell? → **The two electrodes (anode and cathode).**

What is the result of the difference in electrical potential between the electrodes? → **Electromotive force (EMF)**

Electrons generated at the anode (oxidation) are pushed to the cathode (reduction) by electromotive force (EMF).

What is the electromotive force called, and what is its unit? → **The cell potential, the unit is volts.**

What does it mean that the electric potential difference of a voltaic cell is an indication of the energy available to move electrons from the anode to the cathode? → **The higher the potential difference, the more energy is available to move the electrons.**



Roller Coaster

It has gravitational potential energy (PE) at the highest point and converted into kinetic energy (KE) at the lowest point.

The greater the difference in height between the top of the track and the bottom, the higher the kinetic energy.

Calculation of electrochemical cell potentials

Reduction potential (E): the tendency of a material to gain electrons.

Why can't the reduction potential of an electrode be determined directly? → **Because the reduction half-reaction must be combined with the redox half-reaction.**

What is the potential output when the two halves of the reaction are coupled together, and what is its unit? → **The electric potential difference, and its unit is volts.**

Zinc-copper cell

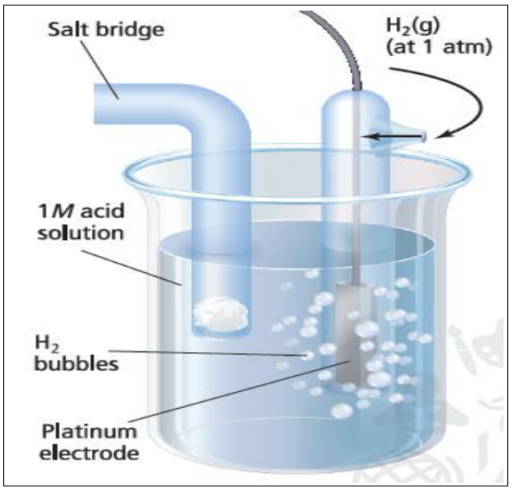
- zinc atoms (Zn) are oxidized
  - zinc has a lower tendency to gain electrons
    - zinc has lower reducing potential
- copper ions (Cu^{2+}) are reduced
  - copper has a greater tendency to gain electrons
    - copper has higher reduction potential

The redox reaction occurs spontaneously only when electrons flow from zinc to copper.

Standard Hydrogen Electrode

Standard Hydrogen Electrode

- It consists of a small plate of platinum immersed in a solution of hydrochloric acid at a concentration of 1M
  - The standard reduction potential of a standard hydrogen electrode is zero.
- Hydrogen H_2 gas is pumped into the solution at standard conditions
  - 25 °C
  - 1 atm



The standard hydrogen electrode functions as an "oxidation half-reaction" or "reduction half-reaction" depending on the half-cell connected to it.

$H_{2(g)} \rightarrow 2H^+_{(aq)} + 2e^-$ (oxidation) $\quad E^\circ_{H_2} = 0.0\ V$

$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$ (reduction) $\quad E^\circ_{H_2} = 0.0\ V$

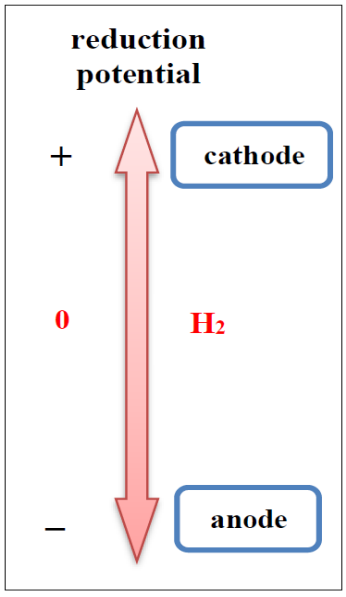
Half-cell potentials

To determine the reduction potential of the elements, the semi-cells were connected with the hydrogen half-cell.

The half of the cell that is more reducible than hydrogen will reduce and have a positive reduction potential.

The half of the cell that is less reducible than hydrogen will oxidize and have a negative reduction potential.

The electrode with higher reduction potential is the cathode, and the electrode with lower reduction potential is the anode.



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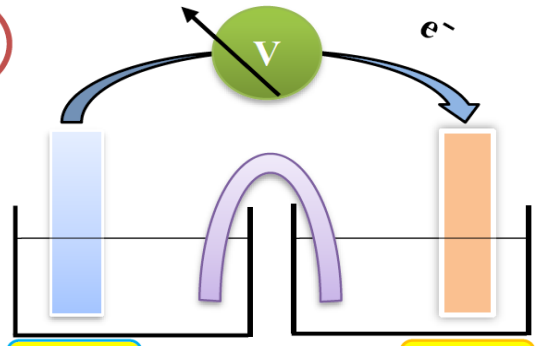


H_2 | X

anode | cathode

less | more

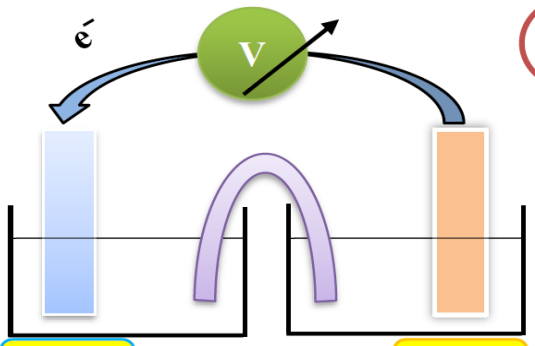
reduction potential

0 | +

$H_{2(g)} \rightarrow 2H^+_{(aq)} + 2e^-$

$H_2 \mid H^+ \parallel X^+ \mid X$

2



H_2 | Y

cathode | anode

more | less

reduction potential

0 | –

$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$

$Y \mid Y^+ \parallel H^+ \mid H_2$

The electrode is measured under standard conditions of 1atm, 25°C and dipped in a 1M solution.

The reduction potential at standard conditions is denoted by the symbol E°

The half-reactions are recorded in the table as reduction reactions.

Standard Reduction Potentials

not required to save

Half-Reaction	E^0 (V)	Half-Reaction	E^0 (V)
$Li^+ + e^- \rightleftarrows Li$	−3.0401	$Cu^{2+} + e^- \rightleftarrows Cu^+$	+0.153
$Ca^{2+} + 2e^- \rightleftarrows Ca$	−2.868	$Cu^{2+} + 2e^- \rightleftarrows Cu$	+0.3419
$Na^+ + e^- \rightleftarrows Na$	−2.71	$O_2 + 2H_2O + 4e^- \rightleftarrows 4OH^-$	+0.401
$Mg^{2+} + 2e^- \rightleftarrows Mg$	−2.372	$I_2 + 2e^- \rightleftarrows 2I^-$	+0.5355
$Be^{2+} + 2e^- \rightleftarrows Be$	−1.847	$Fe^{3+} + e^- \rightleftarrows Fe^{2+}$	+0.771
$Al^{3+} + 3e^- \rightleftarrows Al$	−1.662	$NO_3^- + 2H^+ + e^- \rightleftarrows NO_2 + H_2O$	+0.775
$Mn^{2+} + 2e^- \rightleftarrows Mn$	−1.185	$Hg_2^{2+} + 2e^- \rightleftarrows 2Hg$	+0.7973
$Cr^{2+} + 2e^- \rightleftarrows Cr$	−0.913	$Ag^+ + e^- \rightleftarrows Ag$	+0.7996
$2H_2O + 2e^- \rightleftarrows H_2 + 2OH^-$	−0.8277	$Hg^{2+} + 2e^- \rightleftarrows Hg$	+0.851
$Zn^{2+} + 2e^- \rightleftarrows Zn$	−0.7618	$2Hg^{2+} + 2e^- \rightleftarrows Hg_2^{2+}$	+0.920
$Cr^{3+} + 3e^- \rightleftarrows Cr$	−0.744	$NO_3^- + 4H^+ + 3e^- \rightleftarrows NO + 2H_2O$	+0.957
$S + 2e^- \rightleftarrows S^{2-}$	−0.47627	$Br_2(l) + 2e^- \rightleftarrows 2Br^-$	+1.066
$Fe^{2+} + 2e^- \rightleftarrows Fe$	−0.447	$Pt^{2+} + 2e^- \rightleftarrows Pt$	+1.18
$Cd^{2+} + 2e^- \rightleftarrows Cd$	−0.4030	$O_2 + 4H^+ + 4e^- \rightleftarrows 2H_2O$	+1.229
$PbI_2 + 2e^- \rightleftarrows Pb + 2I^-$	−0.365	$Cl_2 + 2e^- \rightleftarrows 2Cl^-$	+1.35827
$PbSO_4 + 2e^- \rightleftarrows Pb + SO_4^{2-}$	−0.3588	$Au^{3+} + 3e^- \rightleftarrows Au$	+1.498
$Co^{2+} + 2e^- \rightleftarrows Co$	−0.28	$MnO_4^- + 8H^+ + 5e^- \rightleftarrows Mn^{2+} + 4H_2O$	+1.507
$Ni^{2+} + 2e^- \rightleftarrows Ni$	−0.257	$Au^+ + e^- \rightleftarrows Au$	+1.692
$Sn^{2+} + 2e^- \rightleftarrows Sn$	−0.1375	$H_2O_2 + 2H^+ + 2e^- \rightleftarrows 2H_2O$	+1.776
$Pb^{2+} + 2e^- \rightleftarrows Pb$	−0.1262	$Co^{3+} + e^- \rightleftarrows Co^{2+}$	+1.92
$Fe^{3+} + 3e^- \rightleftarrows Fe$	−0.037	$S_2O_8^{2-} + 2e^- \rightleftarrows 2SO_4^{2-}$	+2.010
$2H^+ + 2e^- \rightleftarrows H_2$	**0.0000**	$F_2 + 2e^- \rightleftarrows 2F^-$	+2.866

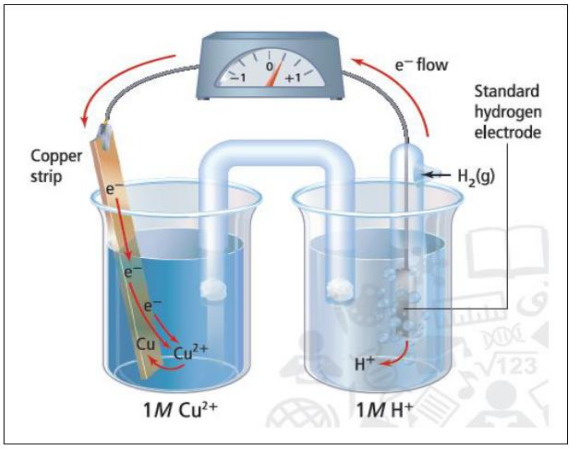
Copper-Hydrogen cell

Electrons flow from the hydrogen electrode to the copper electrode.

E° value measured in voltmeters +0.342 V

Cu^{2+} ions at the copper electrode gain electrons more easily than the H^+ ions at the standard hydrogen electrode.

Copper ions have a greater reduction potential than hydrogen ions.



Copper represents the cathode.

Hydrogen represents the anode

The hydrogen reduction potential is zero, so the copper reduction potential is positive.

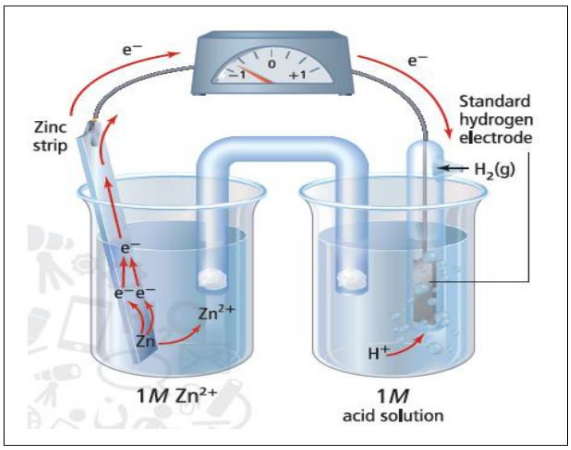
Zinc-Hydrogen cell

Electrons flow from the zinc electrode to the hydrogen electrode.

E° value measured in voltmeters –0.762 V

H^+ ions at the standard hydrogen electrode gain electrons more easily than Zn^{2+} ions at the zinc electrode.

Hydrogen ions have a greater reduction potential than zinc ions.



Zinc represents the anode.

Hydrogen represents the cathode

The hydrogen reduction potential is zero, so the zinc reduction potential is negative.

Electrochemical cell potential calculation

$$E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$$

quantity		unit
symbol	meaning	
E°_{cell}	standard overall potential of the cell	volt (V)
$E^{\circ}_{cathode}$	standard half-cell reduction potential of the reduction	volt (V)
E°_{anode}	standard half-cell reduction potential of the oxidation	volt (V)

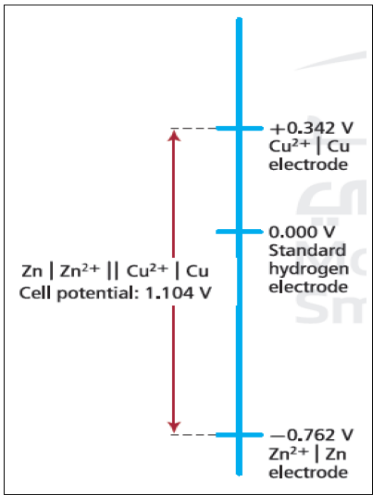
The value of the voltaic cell potential is always positive, so that electrons flow spontaneously from the anode to the cathode.

Example

Calculating the potential of a copper-zinc cell

$$E^{\circ}_{cell} = E^{\circ}_{Cu^{2+}|Cu} - E^{\circ}_{Zn^{2+}|Zn}$$

$$= +0.342 - (-0.762) = +1.104\ V$$



How to determine the anode and the cathode?

1. Direction of movement of electrons.
2. Half-reactions.
3. Overall reactions.
4. Cell notation.
5. Reduction potential.

Study the following half-reduction reactions and answer questions (16-18)

$I_{2(s)} + 2e^- \rightarrow 2I^-_{(aq)}$ $\quad E^\circ_{I_2|I^-} = +0.536\ V$

$Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$ $\quad E^\circ_{Fe^{2+}|Fe} = -0.447\ V$

16) What is the standard cell potential?

a. –0.983 V

b. –0.089 V

c. +0.983 V

d. +0.089 V

17) What is the overall cell reaction?

a. $Fe^{2+}_{(aq)} + I_{2(s)} \rightarrow Fe_{(s)} + 2I^-_{(aq)}$

b. $Fe^{2+}_{(aq)} + 2I^-_{(aq)} \rightarrow Fe_{(s)} + I_{2(s)}$

c. $Fe_{(s)} + 2I^-_{(aq)} \rightarrow Fe^{2+}_{(aq)} + I_{2(s)}$

d. $Fe_{(s)} + I_{2(s)} \rightarrow Fe^{2+}_{(aq)} + 2I^-_{(aq)}$

18) What is the cell notation?

a. $I_2 \mid I^- \parallel Fe \mid Fe^{2+}$

b. $Fe \mid Fe^{2+} \parallel I_2 \mid I^-$

c. $I^- \mid I_2 \parallel Fe^{2+} \mid Fe$

d. $Fe^{2+} \mid Fe \parallel I^- \mid I_2$

19) What is the balanced equation for the spontaneous cell reaction that occurs in the cell with the following reduction half-reactions?

$Mg^{2+}_{(aq)} + 2e^- \rightarrow Mg_{(s)}$ $\quad (E^\circ_{Mg} = -2.37\ V)$

$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$ $\quad (E^\circ_{H_2} = 0.0\ V)$

a. $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$

b. $Mg^{2+} + H_2 \rightarrow Mg + 2H^+$

c. $Mg^{2+} + 2H^+ \rightarrow Mg + H_2$

d. $Mg + H_2 \rightarrow Mg^{2+} + 2H^+$

The battery is constructed using tin and mercury which have the half-reduction reactions:

$Sn^{2+} + 2e^- \rightarrow Sn$ ($E^\circ_{Sn^{2+}|Sn} = -0.138$ V), $Hg^{2+} + 2e^- \rightarrow Hg$ ($E^\circ_{Hg^{2+}|Hg} = +0.851$ V)

Using the values of the reduction potential, answer the questions (20-24).

20) What is oxidized and what is reduced?

a. oxidation: Sn^{2+}, reduction: Hg

b. oxidation: Sn, reduction: Hg^{2+}

c. oxidation: Hg^{2+}, reducing: Sn

d. oxidation: Hg, reducing: Sn^{2+}

21) What are the reactions that takes place at the anode and at the cathode?

choice	anode reaction	cathode reaction
a.	$Hg^{2+} + 2e^- \rightarrow Hg$	$Sn \rightarrow Sn^{2+} + 2e^-$
b.	$Sn \rightarrow Sn^{2+} + 2e^-$	$Hg^{2+} + 2e^- \rightarrow Hg$
c.	$Sn^{2+} + 2e^- \rightarrow Sn$	$Hg \rightarrow Hg^{2+} + 2e^-$
d.	$Hg \rightarrow Hg^{2+} + 2e^-$	$Sn^{2+} + 2e^- \rightarrow Sn$

22) What is the balanced equation for the overall cell reaction?

a. $Hg_{(s)} + Sn^{2+}_{(aq)} \rightarrow Hg^{2+}_{(aq)} + Sn_{(s)}$

b. $Hg_{(s)} + Sn_{(s)} \rightarrow Hg^{2+}_{(aq)} + Sn^{2+}_{(aq)}$

c. $Hg^{2+}_{(aq)} + Sn_{(s)} \rightarrow Hg_{(s)} + Sn^{2+}_{(aq)}$

d. $Hg^{2+}_{(aq)} + Sn^{2+}_{(aq)} \rightarrow Hg_{(s)} + Sn_{(s)}$

23) What is the value of the cell potential (V)?

a. +0.713

b. –0.713

c. +0.989

d. –0.989

24) If there is a sodium sulfate solution in the brine. In which direction do the sulfate ions move?

a. towards the Hg electrode

b. do not move

c. towards the cathode half-cell

d. towards the Sn half-cell

25) What is the balanced chemical equation for a standard cell that has the following notation? $Sn \mid Sn^{2+} \parallel Ag^{+} \mid Ag$

a. $Sn^{2+} + 2Ag^{+} \rightarrow Sn + 2Ag$

b. $Sn + 2Ag \rightarrow Sn^{2+} + 2Ag^{+}$

c. $Sn^{2+} + 2Ag \rightarrow Sn + 2Ag^{+}$

d. $Sn + 2Ag^{+} \rightarrow Sn^{2+} + 2Ag$

26) What does the cell notation when the next half of the cell is attached to the standard hydrogen electrode ($E^{\circ}_{H_2} = 0.0$ V)? $Hg^{2+} \mid Hg \quad E^{\circ}_{Hg^{2+}|Hg} = +0.851$ V

a. $2H^{+} \mid H_2 \parallel Hg \mid Hg^{2+}$

b. $H_2 \mid 2H^{+} \parallel Hg^{2+} \mid Hg$

c. $Hg^{2+} \mid Hg \parallel H_2 \mid 2H^{+}$

d. $Hg \mid Hg^{2+} \parallel 2H^{2+} \mid H_2$

27) What is the value of the cell potential (V) that has the following cell notation?

$S^{2-} \mid S \parallel Au^{+} \mid Au \qquad (E^{\circ}_{S} = -0.47$ V, $E^{\circ}_{Au} = +1.69$ V)

a. +2.16

b. –2.16

c. +1.22

d. –1.22

28) What are the electrodes make up a cell where $E^{\circ} = +0.05$ V?

($Ag^{+} = +0.80$ V, $Hg^{2+} = +0.85$ V, $Cu^{2+} = +0.34$ V, $Mg^{2+} = -2.37$ V)

a. Ag, Cu

b. Mg, Cu

c. Hg, Ag

d. Hg, Mg

Study the following half-reduction reactions and answer questions (29-31)

$Co^{2+}_{(aq)} + 2e^- \rightarrow Co_{(s)}$ $\quad E^\circ_{Co^{2+}|Co} = -0.28\ V$

$Cr^{3+}_{(aq)} + 3e^- \rightarrow Cr_{(s)}$ $\quad E^\circ_{Cr^{3+}|Cr} = -0.744\ V$

29) What is the standard cell potential?

a. –0.464 V

b. –1.024 V

c. +0.464 V

d. +1.024 V

30) What is the overall cell reaction?

a. $2Cr^{3+} + 3Co^{2+} \rightarrow 2Cr + 3Co$

b. $2Cr + 3Co \rightarrow 2Cr^{3+} + 3Co^{2+}$

c. $2Cr^{3+} + 3Co \rightarrow 2Cr + 3Co^{2+}$

d. $2Cr + 3Co^{2+} \rightarrow 2Cr^{3+} + 3Co$

31) What is the cell notation?

a. $Co \mid Co^{2+} \parallel Cr^{3+} \mid Cr$

b. $Cr \mid Cr^{3+} \parallel Co^{2+} \mid Co$

c. $Co^{2+} \mid Co \parallel Cr \mid Cr^{3+}$

d. $Cr^{3+} \mid Cr \parallel Co \mid Co^{2+}$

32) What is the potential of the cell in which the following reaction takes place?

$2Ag^+_{(aq)} + Pb_{(s)} \rightarrow Pb^{2+}_{(aq)} + 2Ag_{(s)}$

$(E^\circ_{Ag} = +0.799\ V,\ E^\circ_{Pb} = -0.126\ V)$

a. +0.925 V

b. –0.925 V

c. +0.673 V

d. –0.673 V

Use the following table to answer questions (33–36)

Selected Standard Reduction Potentials at 25°C, 1 atm, and 1*M* Ion Concentration	
Half-Reaction	**E^0 (V)**
$Mg^{2+} + 2e^- \rightarrow Mg$	−2.372
$Al^{3+} + 3e^- \rightarrow Al$	−1.662
$Pb^{2+} + 2e^- \rightarrow Pb$	−0.1262
$Ag^+ + e^- \rightarrow Ag$	0.7996
$Hg^{2+} + 2e^- \rightarrow Hg$	0.851

33) Which ion is the easiest to reduce?

a. Mg^{2+} b. Al^{3+}

c. Ag^+ d. Hg^{2+}

34) What is the correct cell notation for the cell?

a. $Ag \mid Ag^+ \parallel Al^{3+} \mid Al$

b. $Mg \mid Mg^{2+} \parallel H^+ \mid H_2$

c. $H_2 \mid H^+ \parallel Pb^{2+} \mid Pb$

d. $Pb \mid Pb^{2+} \parallel Al^{3+} \mid Al$

35) What is the standard potential of a cell (V) of a magnesium electrode and a silver electrode each in a 1M solution?

a. +1.572 b. +0.773

c. +3.172 d. +3.971

36) What cell produces a potential of +2.513 V, under a standard conditions?

a. $Al \mid Al^{3+} \parallel Hg^{2+} \mid Hg$

b. $H_2 \mid H^+ \parallel Hg^{2+} \mid Hg$

c. $Mg \mid Mg^{2+} \parallel Al^{3+} \mid Al$

d. $Pb \mid Pb^{2+} \parallel Ag^+ \mid Ag$

37) What is the balanced chemical equation for a standard cell that has the following notation? $I^- \mid I_2 \parallel Fe^{3+} \mid Fe^{2+}$

a. $I_2 + 2Fe^{2+} \rightarrow 2I^- + 2Fe^{3+}$

b. $2I^- + 2Fe^{3+} \rightarrow I_2 + 2Fe^{2+}$

c. $I_2 + 2Fe^{3+} \rightarrow 2I^- + 2Fe^{2+}$

d. $2I^- + 2Fe^{2+} \rightarrow I_2 + 2Fe^{3+}$

38) What does the cell notation when the next half of the cell is attached to the standard hydrogen electrode ($E^{\circ}_{H_2} = 0.0$ V)? $Al^{3+} \mid Al$ $E^{\circ}_{Al^{3+}|Al} = -1.662$ V

a. $Al \mid Al^{3+} \parallel 2H^{2+} \mid H_2$

b. $Al^{3+} \mid Al \parallel H_2 \mid 2H^{+}$

c. $H_2 \mid 2H^{+} \parallel Al^{3+} \mid Al$

d. $2H^{+} \mid H_2 \parallel Al \mid Al^{3+}$

39) What is the value of the cell potential (V) that has the following cell notation?

$H_2 \mid 2H^{+} \parallel 2I^{-} \mid I_2$ $(E^{\circ}_{H_2} = 0.0$ V, $E^{\circ}_{I_2} = +0.53$ V)

a. –1.06 b. +1.06

a. –0.53 b. +0.53

40) Which expression describes a voltaic cell consisting of a cadmium electrode and a copper electrode? $(E^{\circ}_{Cd} = -0.40$ V, $E^{\circ}_{Cu} = +0.34$ V)

a. reduction of cadmium

b. oxidation of copper

c. the positive ions move from the copper half-cell to the cadmium half-cell

d. electrons move from the cadmium half-cell to the copper half-cell

Study the following cell notation: $Al \mid Al^{3+} \parallel Zn^{2+} \mid Zn$, and answer questions (41,42)

41) What happens to the mass of the aluminum electrode, and why?

a. decrease, oxidize b. decrease, reduce

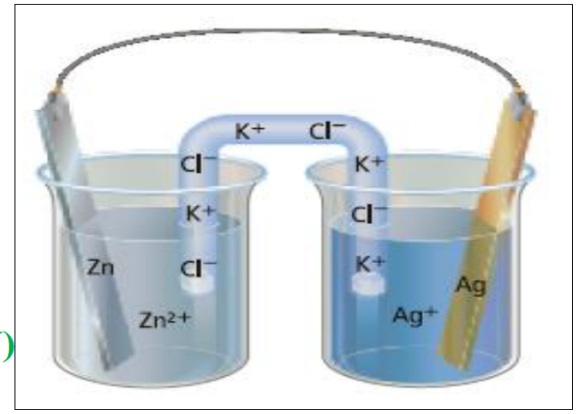
c. increases, oxidizes d. increases, reduce

42) What is the reduction potential for zinc (V) if the reduction potential of aluminum is -1.66 V and the cell potential is +0.9 V?

a. –2.56 b. +2.56

c. –0.76 d. +0.76

The corresponding figure shows a voltaic cell consisting of a strip of zinc in 0.1 M zinc nitrate solution and a strip of silver in 0.1 M silver nitrate solution. Use the plot and values for standard reduction potentials to answer the questions (43,44). ($E^{\circ}_{Ag^+|Ag} = +0.7996$ V, $E^{\circ}_{Zn^+|Zn} = -0.7618$ V)



43) What is the value of the cell potential (V)?

a. +0.0378 b. –0.0378

c. +1.5614 d. –1.5614

44) Which of the following statements describes what happens in the cell?

a. the following reaction: $Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn$ takes place at the zinc electrode

b. electrons move through the salt bridge to the silver side

c. the ions pass through the wire from the silver strip to the zinc strip

d. positive silver ions build up around the silver strip

45) Based on the standard reduction potentials mentioned in the following table, if a silver electrode is connected to a chromium electrode in a voltaic cell, which electrode will be oxidized and which will be reduced?

Standard Reduction Potentials at 25°C, 1 atm, and 1*M* Solution	
$Ag^+ + e^- \rightarrow Ag$	0.7996
$Cr^{3+} + 3e^- \rightarrow Cr$	−0.744

a. oxidation: Cr, reduction: Ag^+

b. oxidation: Ag, reduction: Cr^{3+}

c. oxidation: Cr^{3+}, reduction: Ag

d. oxidation: Ag^+, reduction: Cr

29	c	30	d	31	b	32	a	33	d
34	b	35	c	36	a	37	b	38	a
39	d	40	d	41	a	42	c	43	c
44	d	45	a						

Use standard reduction potential

How do electrons flow in a voltaic cell? → **From half of the cell with low reduction potential (anode) to half of cell with high reduction potential (cathode).**

What results from the flow of electrons in the voltaic cell? → **It causes a positive potential to the cell.**

determine the spontaneity of a reaction

if the cell potential is positive

the reaction is sponteneous

sponteneous

reactants ⇄ products

non-sponteneous

if the cell potential is negative

the reaction is non-sponteneous

non-sponteneous

reactants ⇄ products

sponteneous

How to determine the anode and the cathode?

half-reactions

oxidation	reduction
$X \rightarrow X^+ + e^-$	$Y^+ + e^- \rightarrow Y$
X: anode	Y^+: cathode

half-reactions with reduction potential

$X^+ + e^- \rightarrow X$ $\quad E^\circ_X = \ldots\ldots V$

$Y^+ + e^- \rightarrow Y$ $\quad E^\circ_Y = \ldots\ldots V$

$E^\circ_X < E^\circ_Y$

X: anode $\quad$ Y^+: cathode

overall reaction

$X + Y^+ \rightarrow X^+ + Y$

X: anode $\quad$ Y^+: cathode

cell notation

$X \mid X^+ \parallel Y^+ \mid Y$

X: anode $\quad$ Y^+: cathode

46) What is the cell potential for the following balanced redox reaction, and is the reaction spontaneous or not?

$Sn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Sn^{2+}_{(aq)} + Cu_{(s)}$ $(E^{\circ}_{Sn} = -0.1375\ V,\ E^{\circ}_{Cu} = +0.3419\ V)$

choice	E°_{cell} (V)	spontaneity
a.	+0.4794	non-spontaneous
b.	+0.4794	spontaneous
c.	–0.4794	non-spontaneous
d.	–0.4794	spontaneous

47) What is the cell potential for the following balanced redox reaction, and is the reaction spontaneous or not?

$2SO_4^{2-}{}_{(aq)} + Co^{2+}_{(aq)} \rightarrow Co_{(s)} + S_2O_8^{2-}{}_{(aq)}$ $(E^{\circ}_{SO_4\ 2-} = +2.01\ V,\ E^{\circ}_{Co} = -0.28\ V)$

choice	E°_{cell} (V)	spontaneity
a.	+2.29 V	spontaneous
b.	+2.29 V	non-spontaneous
c.	–2.29 V	spontaneous
d.	–2.29 V	non-spontaneous

48) What is the cell potential for the following balanced redox reaction, and is the reaction spontaneous or not?

$Al \mid Al^{3+} \parallel Hg^{2+} \mid Hg$ $(E^{\circ}_{Al} = -1.662\ V,\ E^{\circ}_{Hg} = +0.851\ V)$

choice	E°_{cell} (V)	spontaneity
a.	+2.513	spontaneous
b.	+2.513	non-spontaneous
c.	–2.513	spontaneous
d.	–2.513	non-spontaneous

49) What is the cell potential for the following balanced redox reaction, and is the reaction spontaneous or not?

$Mg_{(s)} + Pb^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Pb_{(s)}$ $(E^{\circ}_{Mg} = -2.372\ V,\ E^{\circ}_{Pb} = -0.1262\ V)$

choice	E°_{cell} (V)	spontaneity
a.	–2.4982	non-spontaneous
b.	+2.4982	spontaneous
c.	–2.2458	non-spontaneous
d.	+2.2458	spontaneous

50) What is the cell potential for the following balanced redox reaction, and is the reaction spontaneous or not?

$2Al^{3+}_{(aq)} + 3Cu_{(s)} \rightarrow 3Cu^{2+}_{(aq)} + 2Al_{(s)}$ $(E^{\circ}_{Al} = -1.662\ V,\ E^{\circ}_{Cu} = +0.341\ V)$

choice	E°_{cell} (V)	spontaneity
a.	–1.321	spontaneous
b.	+1.321	non-spontaneous
c.	–2.003	non-spontaneous
d.	+2.003	spontaneous

51) What is the cell potential for the following balanced redox reaction, and is the reaction spontaneous or not?

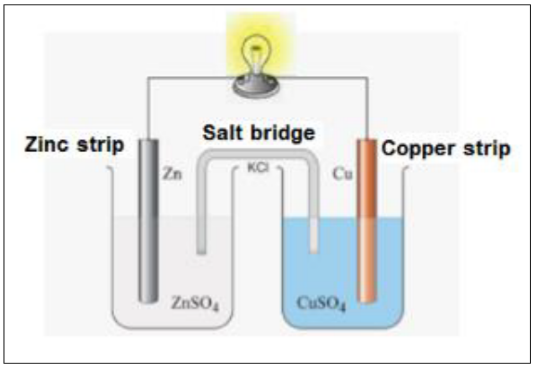
$Hg^{2+}_{(aq)} + 2Cu^{+}_{(aq)} \rightarrow 2Cu^{2+}_{(aq)} + Hg_{(l)}$ $(E^{\circ}_{Hg} = +0.851\ V,\ E^{\circ}_{Cu} = +0.153\ V)$

choice	E°_{cell} (V)	spontaneity
a.	–0.698	non-spontaneous
b.	+0.698	spontaneous
c.	+1.004	spontaneous
d.	–1.004	non-spontaneous

49	d	50	c	51	b				

2021-2022 Exam Questions

1. **In the cell shown in the figure below, which of the following statements is true?**



a. uses electrical energy to cause a chemical reaction

b. positive zinc ions build up around the zinc electrode

c. electrons move through the salt bridge to the copper side

d. the wire serves as a pathway for electrons to flow from the zinc strip to the copper strip

2. **The following reduction half-reactions represent the half-cells of a voltaic cell, which of the following is correct?**

	anode	cathode	Electrochemical cell potential (E°)	Reaction type
A	Pt	Co	+0.9 V	spontaneous
B	Co	Pt	+1.46 V	spontaneous
C	Pt	Co	–0.9 V	nonspontaneous
D	Co	Pt	–1.46 V	nonspontaneous

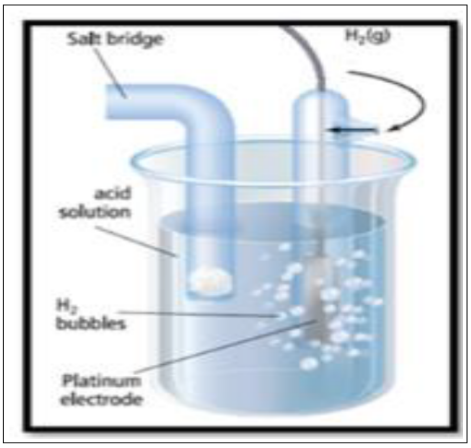
a. B

b. D

c. C

d. A

3. **Which of the following is true of the standard hydrogen electrode shown in the figure below?**



a. the reduction potential of this electrode is 0.000 V at all conditions

b. it is the anode in all the voltaic cells in which it is connected

c. the reduction potential of this electrode is 0.000 V when the concentration of the acid solution is 1 M and hydrogen gas is pumped under a pressure of 1 atm and the temperature is kept constant at 25 °C

d. it is the cathode in all the voltaic cells in which it is connected

1	d	2	a	3	c				

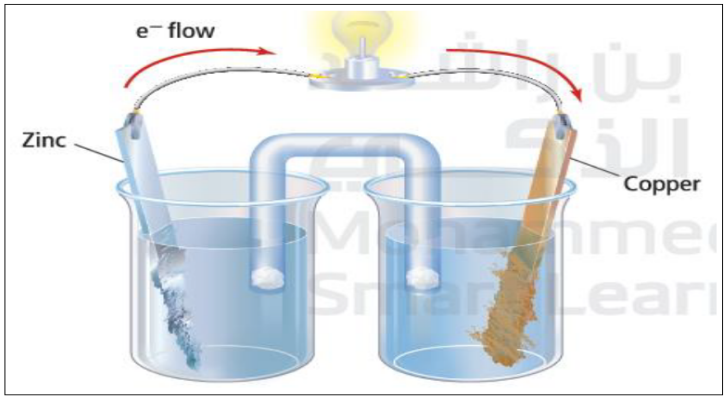
Sec. (3) Electrolysis

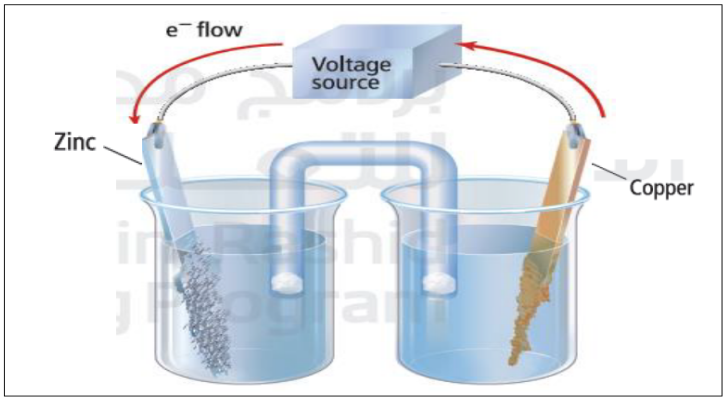
What is produced by the Redox reaction in a voltaic cell?

Electrons flow through the external circuit from the anode electrode to the cathode electrode.

When does the reaction stop in the voltaic cell shown in the corresponding figure?

When the zinc electrode is consumed.





How can the cell renewal?

When a current is applied in the opposite direction using an external (voltage) power supply.

When the voltage source is left long enough, the cell will return to nearly its original power.

voltaic cell

sponteneous

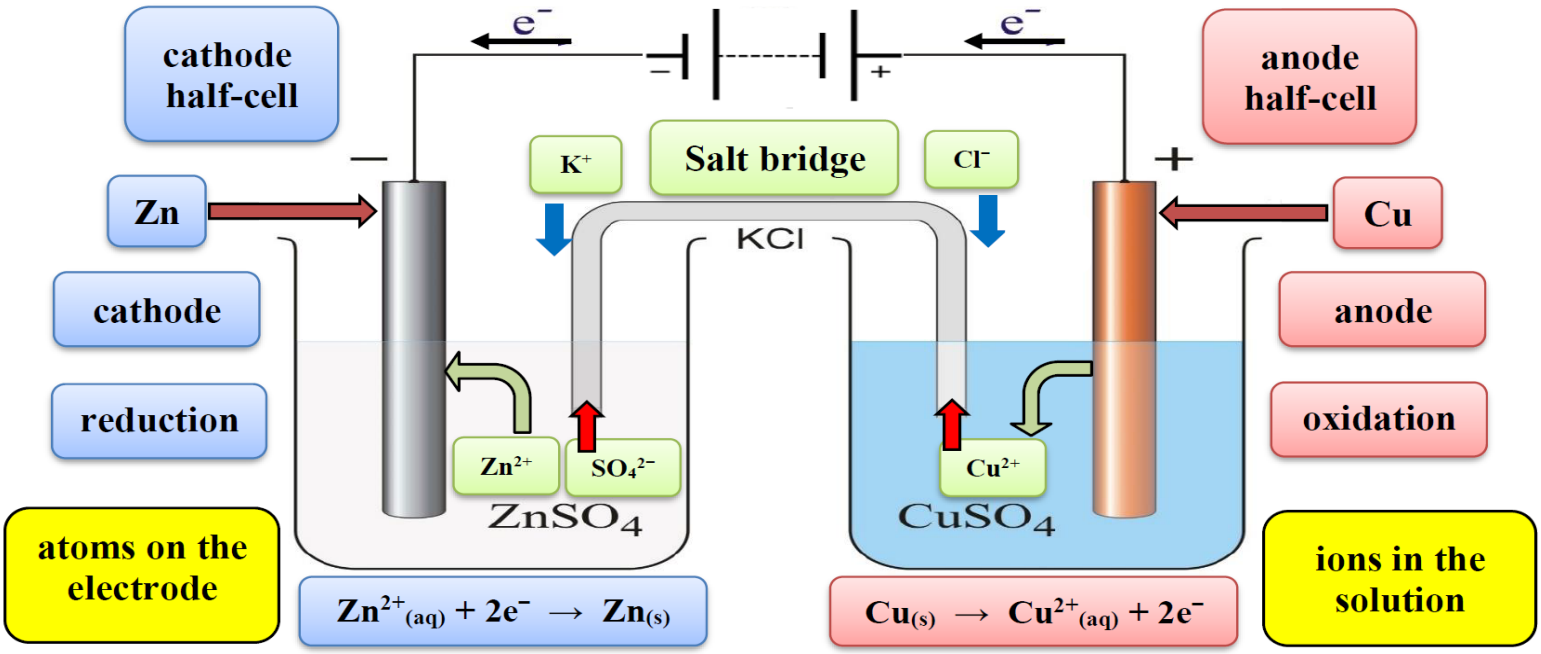
$Zn + Cu^{2+} \rightleftarrows Zn^{2+} + Cu$

non-sponteneous

electrolytic cell

Electrolysis: The use of electrical energy to get a chemical reaction.

Electrolytic Cell: The electrochemical cell in which electrolysis takes place.



The external energy source in the electrolytic cell reverses the electrode reactions.

Comparison	voltaic cell	electrolytic cell
energy transformations	from chemical to electrical	from electrical to chemical
external energy source	no	yes
spontaneity	sponteneous	non-sponteneous
cell potential sign	positive	negative
anode reaction	oxidation	
cathode reaction	reduction	
movement of electrons	from anode to cathode	

52) Which of the following describes electrolytic cells?

a. a spontaneous cell

b. energy is converted from chemical to electrical

c. has a positive potential cell

d. the oxidation process occurs at the anode electrode

53) The hydrolysis cell has the following reaction: $2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$

The fuel cell has the reaction: $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$ $E^\circ_{cell} = +1.229$ V

What is the value of E°_{cell} of the hydrolysis cell (V)?

a. +1.229 b. –1.229 c. 2.458 d. –2.458

54) Which of the following statements describes electrolytic cells?

a. the reduction process occurs at the anode

b. corrosion occurs to the cathode electrode

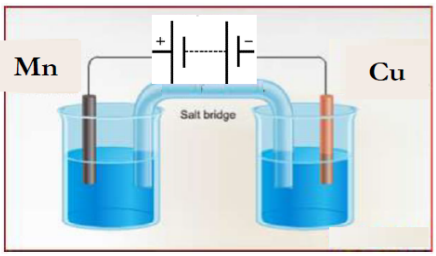
c. the battery reverses the redox reactions

d. the electrons move from the cathode to the anode through the wire

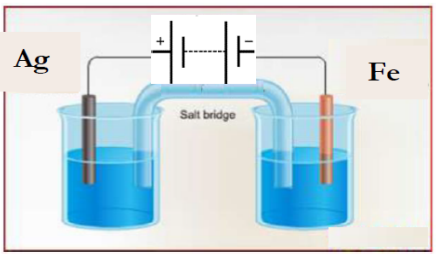
55) Depending on the values of the reduction potentials in the table, which of the following cells represent an electrolytic cell?

Electrode	Ag^+/Ag	Cu^{2+}/Cu	Fe^{2+}/Fe	Mn^{2+}/Mn
E° (V)	+0.799	+0.342	–0.447	–1.185

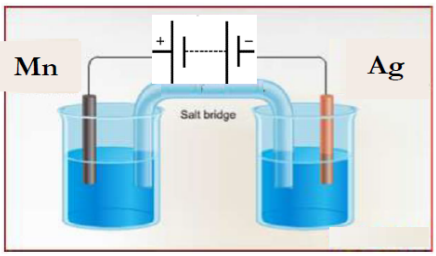
a.



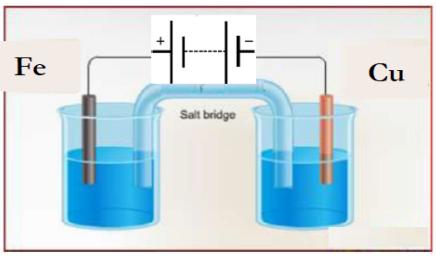
b.



c.



d.



Electrolysis of molten NaCl

Molten sodium chloride can be decomposed into sodium metal and chlorine gas in a cell called Down's cell.

anode	substance: carbon	reaction: $2Cl^-_{(l)} \rightarrow Cl_{2(g)} + 2e^-$
cathode	substance: iron	reaction: $Na^+_{(l)} + e^- \rightarrow Na_{(l)}$
overall reaction: $2Na^+_{(l)} + 2Cl^-_{(l)} \rightarrow 2Na_{(l)} + Cl_{2(g)}$		
electrolyte	molten NaCl	

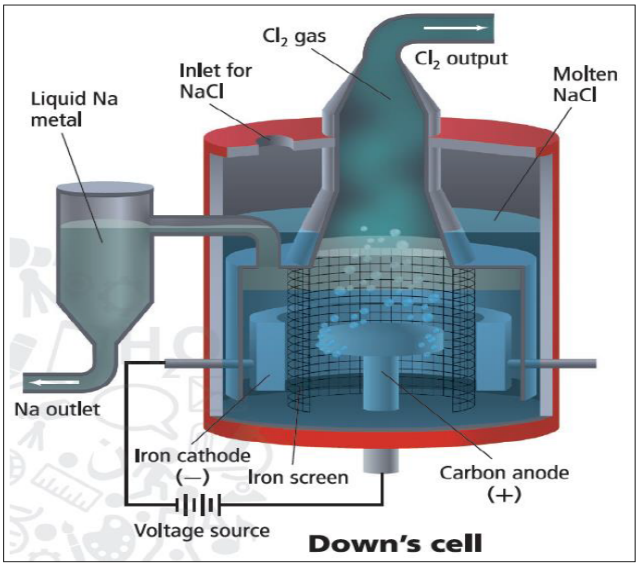
The electrical source provides the electrons which are used to reduce the sodium ions at the cathode.

When electrons are lost from the anode, and the chloride ions at the anode oxidized to form chlorine gas.

Why is NaCl used as molten in Down's cell?

⇩

Because ionic compounds conduct current when they are in solution or molten form where the ions can move.



uses of chlorine

- produce of paper, plastics, pesticides, textiles, dyes and paints
- produce of household detergents such as household bleach
- purified water for drinking and swimming

uses of sodium

- produce many salts used in foods in its ionic compounds
- produce of outdoor sodium vapor lamps
- coolant in nuclear reactions in its pure form

Electrolysis of brine

Decomposition of brine (sodium chloride solution) is a typical electrolytic cell.

anode	reaction: $2Cl^-_{(l)} \rightarrow Cl_{2(g)} + 2e^-$
cathode	reaction: $2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$

overall reaction: $2H_2O_{(l)} + 2NaCl_{(aq)} \rightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{(aq)}$

Which of the two reactions takes place at the anode? why?

$2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$

$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$

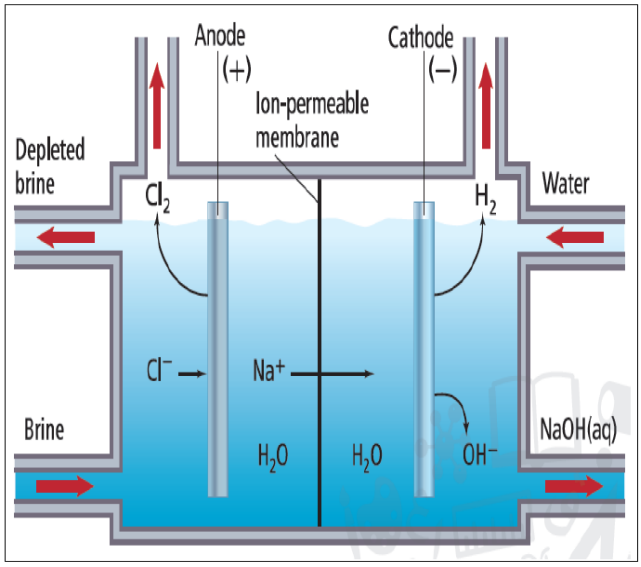
$2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$ because the desired product is chlorine gas, Cl_2, the concentration of chloride ions, Cl^-, is kept high

Which of the two reactions takes place at the cathode? why?

$Na^+_{(aq)} + e^- \rightarrow Na_{(s)}$

$2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$

$2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$ because water is easier to reduce



The materials resulting from the total reaction (H_2 gas, Cl_2 gas, and NaOH solution) are important commercial materials.

Chlorine gas is used in the manufacture of PVC products such as water pipes.

56) What is the anode reaction in a Down's cell?

a. $2Cl^-_{(l)} \rightarrow Cl_{2(g)} + 2e^-$

b. $Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(l)}$

c. $Na_{(l)} \rightarrow Na^+_{(l)} + e^-$

d. $Na^+_{(l)} + e^- \rightarrow Na_{(l)}$

57) What is the reaction of the cathode in a NaCl solution electrolytic cell?

a. $Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}$

b. $O_{2(g)} + 4H^+_{(aq)} + 4e^- \rightarrow 2H_2O_{(l)}$

c. $Na^+_{(aq)} + e^- \rightarrow Na_{(s)}$

d. $2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$

58) Which of the following substances does not result from the overall reaction of the electrolysis of a NaCl solution?

a. $H_{2(g)}$

b. $Cl_{2(g)}$

c. $Na_{(s)}$

d. $NaOH_{(aq)}$

59) Why is NaCl used as molten or solution in electrolytic cells?

①	②	③
ionic compounds in the solid state conduct electricity	molten NaCl contains free moving ions	NaCl solution in which the Na^+ and Cl^- ions move freely

a. ① and ② only

b. ② and ③ only

c. ①, ② and ③

d. ① and ③ only

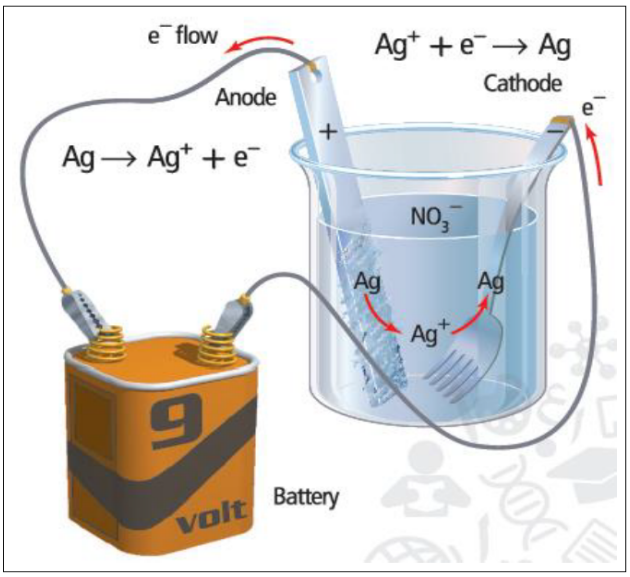
Electroplating

Objects are electroplated by applying a thin and uniform layer of metal.

anode	material to be painted with	connect to the positive of the battery
cathode	material to be painted	connect to the negative of the battery
electrolyte	solution of the material to be painted with.	

Electroplating of an item with silver

anode	silver
anode reaction	$Ag \rightarrow Ag^+ + e^-$
cathode	item to be painted
cathode reaction	$Ag^+ + e^- \rightarrow Ag$
electrolyte	$AgNO_{3(aq)}$

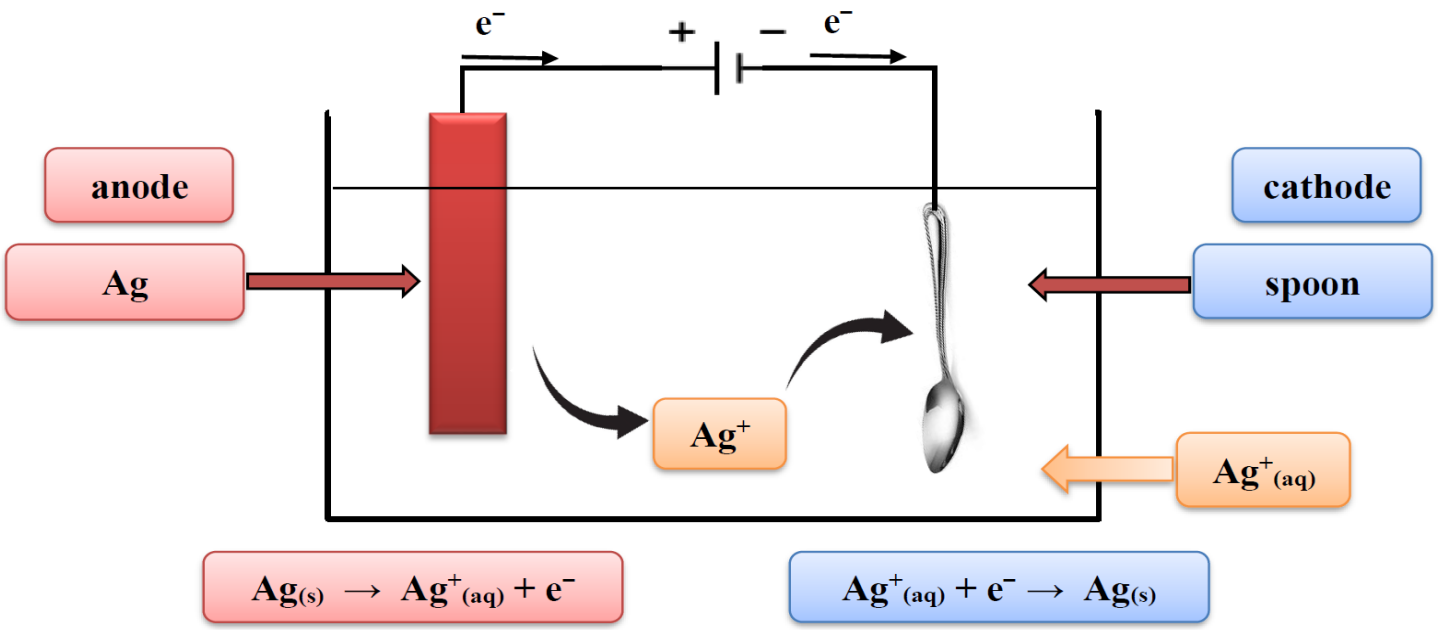


**Silver atoms are transferred from the anode to the cathode,
the silver electrode is eroded and the body is coated with a layer of silver.**

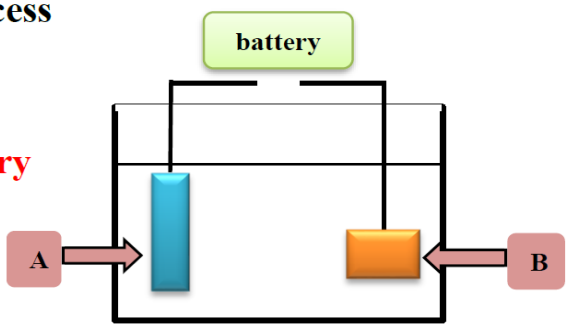
The current passing through the cell must be carefully controlled to obtain a smooth and even layer of metal in the coating.

Some cars are electroplated with nickel and then chromium to make them more rust-resistant.

Electroplating of metallic spoon with silver



60) Which of the following statements describes the process of electroplating object B with metal A?



a. B is connected to the positive electrode of the battery

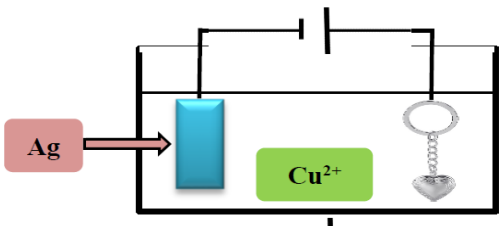
b. cathode reaction: $A \rightarrow A^{2+} + 2e^-$

c. the solution used as the electrolyte contains B ions

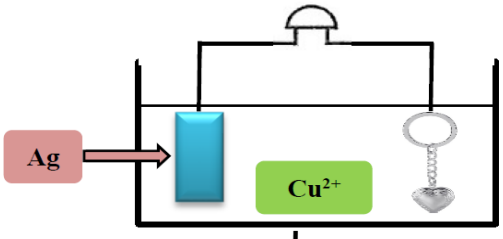
d. mass of A decreases when the plating process is completed

61) Which of the following cells represents an electrolytic cell in which a metal medal is electroplated with silver metal?

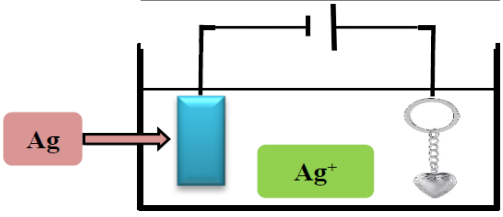
a.



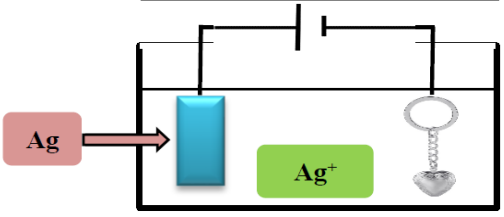
b.



c.



d.



2021-2022 Exam Questions

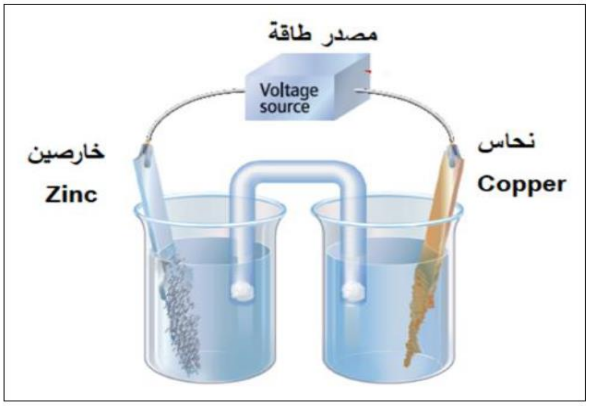
1. **Which of the following is true of the electrochemical cell in the figure below?**

   a. converts chemical energy to electrical energy

   b. non-spontaneous oxidation and reduction reactions occur

   c. spontaneous oxidation and reduction reactions occur

   d. the reaction continues until the zinc strip is used up, then the reaction stops



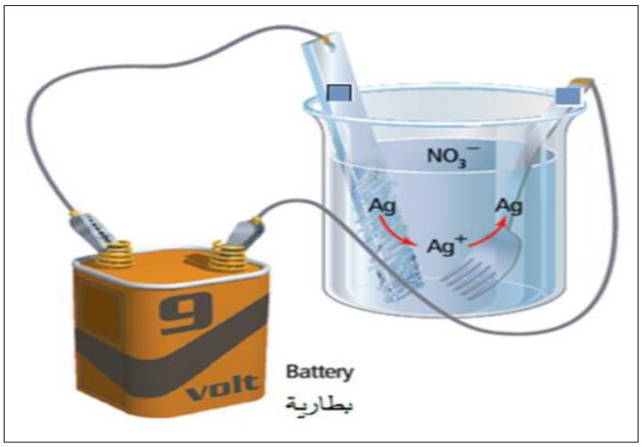
2. **Objects are electroplated when a uniform coating is deposited as a protective or decorative layer, as shown in the figure below. Which of the following is true?**

   a. the metal used for coating is the cathode

   b. silver is oxidized to silver ions at the cathode

   c. silver is oxidized to silver ions at the anode

   d. the object to be plated is the anode



1	b	2	c						