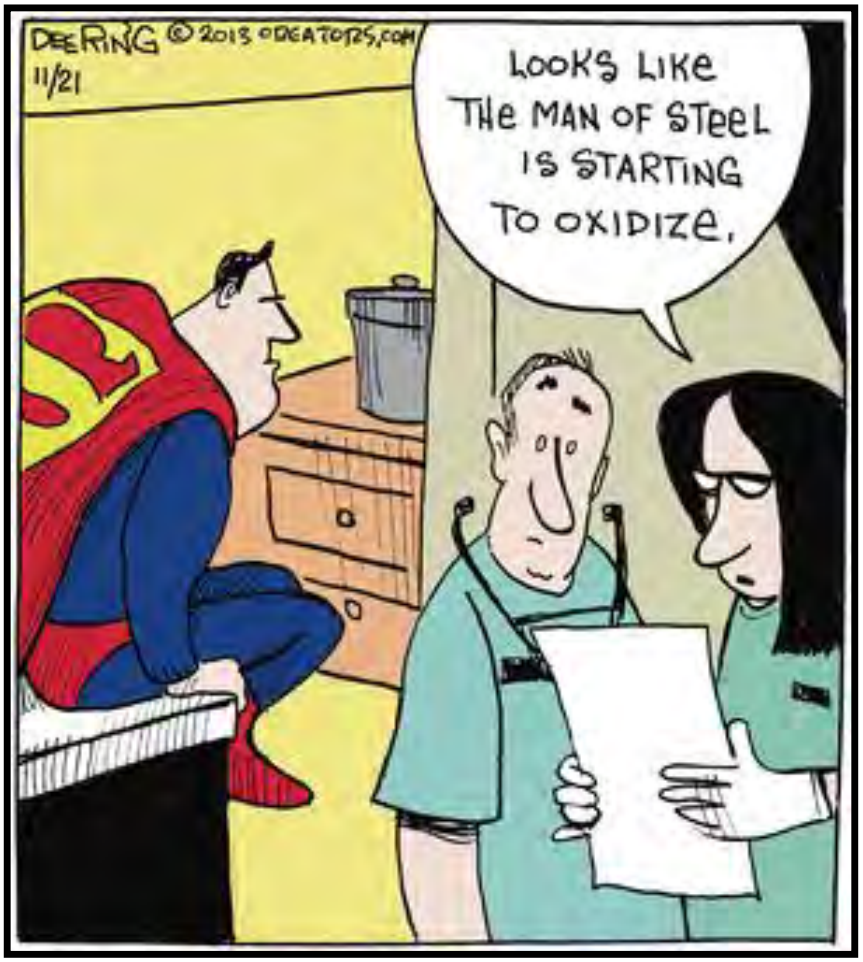
Regents Chemistry: Dr. Shanzer

Practice Packet

Chapter 14: Oxidation Reduction & Electrochemistry



Chapter 14: Electrochemistry

Anode – the electrode at which oxidation occurs (negative)

Cathode – The electrode at which reaction occurs (positive)

Electrode – a conductor in a circuit that carries electrons to or from a substance other than a metal

Electrochemical Cell – any device that converts chemical energy into electrical energy or electrical energy into chemical energy

Electrolytic Cell – a cell that uses electricity from an outside source to force a nonsponstanous redox reaction to occur; Examples: recharging batteries, electroplating

Electroplating – an electrolytic process that involved oxidizing a source metal into a solution with the use of an external power source and then reducing the metal ion onto a metallic object that is to be plated ("silver plated")

Half Reaction – a reaction that describes the change in oxidation number and subsequent gain or loss of electrons that occurs during oxidation or reduction

Oxidation – a process that involves complete or partial loss of electrons or gain of oxygen; results in an increase in the oxidation number

Oxidizing Agent – the species that is reduced and therefore removes the electrons from the species that was oxidized

Oxidation Number – the charge of an ion

Redox reaction – another name for an oxidation-reduction reaction; a reaction that involves the transfer of electrons between reactants

Reduction – a process that involves a complete or partial gain of electrons or the loss of oxygen; it results in a decrease in the oxidation number

Reducing Agent – the species that is oxidized and therefore gives electrons to the species that was reduced.

Salt Bridge – a tube containing a strong electrolyte used to separate the half-cells in a voltaic cell; it allows the passage of ions form one half-cell to the other but prevents the solutions from mixing

Species – the symbol and charge of an element or ion in a redox reaction

Voltaic Cell – an electrochemical cell that produces electric current as a result of a spontaneous redox reaction, used to make batteries. Consists of two half-cells connected by a salt bridge and two electrodes that connect to a load (device) that uses the electricity produced by the cell.

Assigning Oxidation Numbers

Chemistry 200
Video Lesson 14.1

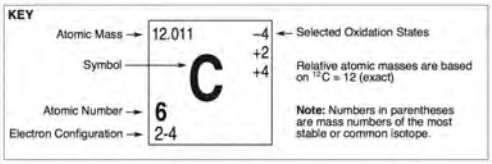
Objective:

How do we assign atoms the correct oxidation number?

Oxidation & Reduction (Redox)

Oxidation number

- a number assigned to keep track of electron(s) gained or lost in a redox rxns.



Rules for Assigning Oxidation Numbers (states)

1. Free elements are assigned an oxidation state of zero

   The hydrogen in H_2, the sodium in Na, and the sulfur in S_8 all have oxidation numbers of **zero.**

2. The Group 1 metals (Alkali Metals) in compounds always have an oxidation state of +1
3. The Group 2 metals (Alkali Earth Metals) in compounds always have an oxidation state of +2
4. The oxidation state for any simple one-atom ion(monoatomic) is equal to its charge.

In $MgCl_2$, the *Mg^{2+} ion* has an oxidation number of +2. Each of the *Cl^- ions* has an oxidation number of *-1*.

In $FeCl_2$, the *Fe^{2+} ion* has an oxidation number of +2, while in $FeCl_3$ the *Fe^{3+} ion* has an oxidation number of +3.

5. Fluorine in compounds is always assigned an oxidation state of -1.
6. Hydrogen in compounds is always assigned an oxidation number of +1. (HCl). If combined w/ a metal, Hydrogen has an oxidation # of -1. (LiH)

7. Oxygen in compounds is assigned an oxidation state of -2.

When combined w/ Fluorine, Oxygen has an oxidation state of +2.

8. The sum of the oxidation states of all the atoms in a species must be equal to the net charge on the species.

If the species is neutral, the sum of the oxidation states is zero.

If the species has a charge, the sum is equal to that charge.

Determine the oxidation number of sulfur in H_2SO_4

$\frac{(+2)}{+1}\ \frac{+6}{\boxed{+6}}\ \frac{(-8)}{-2} = 0$

H_2SO_4

Determine the oxidation number of nitrogen in NO_3^{-1}

$\frac{+5}{\boxed{+5}}\ \frac{(-6)}{-2} = -1$

NO_3^{-1}

Redox Reactions

Chemistry 200
Video Lesson 14.2

Objective:

How do we identify substances being oxidized or reduced and the oxidizing and reducing agents.

How do we use this information to create half reactions ?

Redox

Redox reaction

- a chemical rxn where electron(s) are transferred from one atom to another

Oxidation

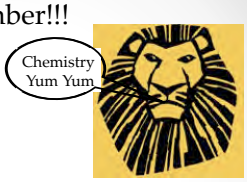
- loss of electron(s) by an atom or ion & an increase in oxidation # ***($Ca \rightarrow Ca^{+2} + 2e^-$)***

Reduction

- gain of electron(s) by an atom or ion & a decrease in oxidation # ***($5e^- + Mn^{+7} \rightarrow Mn^{+2}$)***

*****Oxidation & Reduction occur simultaneously, they cannot occur separately*****

Easy way to remember!!!



LEO says GER

LEO = Loss of Electrons is Oxidation

GER = Gain of Electrons is Reduction

Oxidizing Agent

- the substance that causes the ***oxidation*** of other substances. They accept electron(s) easily & therefore are reduced

ex$_1$: $5e^- + Mn^{+7} \rightarrow Mn^{+2}$

(gain of electrons, oxidation # reduced)

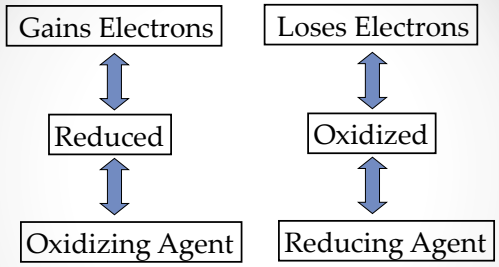
Reducing Agent

- the substance that allows another substance to be ***reduced.*** They lose electron(s) & therefore are oxidized

ex: $Ca \rightarrow Ca^{+2} + 2e^-$

(loss of electrons, oxidation # increases)

The species oxidized, the species reduced, the oxidizing agent or reducing agent must always be a reactant!!



Writing Half Reactions

- They show either the oxidation or reduction portion of a redox rxn, including electrons gained or lost

Half rxn for oxidation --> electrons are products
Half rxn for reduction --> electrons are reactants

THERE'S AN EASIER WAY TO REMEMBER!!!

***** Electrons go on the side w/ the more positive oxid. # *****

Creating Half Reactions from a Redox Reaction

1. Assign oxidation numbers to determine the substance oxidized & the substance reduced.
2. Write each half reaction, balance by atom, then by electrons. The ONLY exception is a half rxn the contains a diatomic element.
3. Identify the species oxidized, the species reduced the oxidizing agent & the reducing agent.

+1 +5 -6
+1 +5 -2 +1 -1 0

$KClO_3 \rightarrow KCl + O_2$

Oxid: $2O^{-2} \rightarrow O_2 + 4e^-$ *(reducing agent)*

Red: $6e^- + Cl^{+5} \rightarrow Cl^{-1}$ *(oxidizing agent)*

Electrochemical Cells

Video Lesson 14.4

Objectives

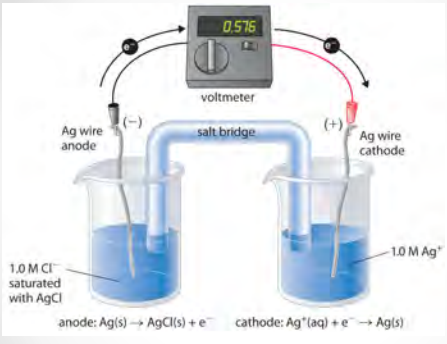
- Name the type of reactions involved in electrochemical processes.
- Describe how a voltaic cell produces electrical energy.

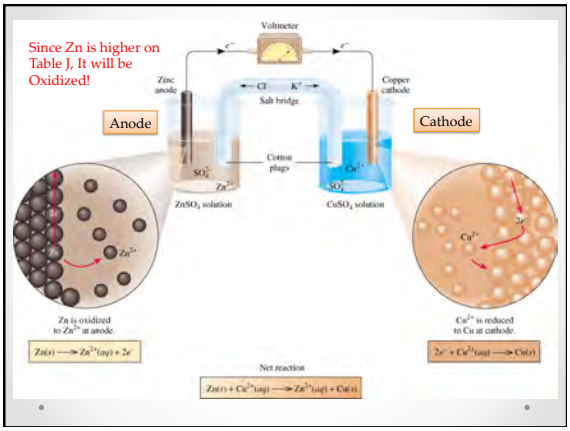
Electrochemical Process

- An electrochemical process is any conversion between chemical and electrical energy (charged particles).
- All electrochemical processes involve redox reactions.
- The two half reactions must be physically separated to be used as a source of electrical energy.

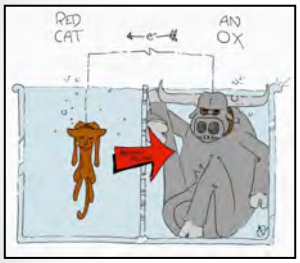
Voltaic Cells

- To obtain a useful current, we separate the oxidizing and reducing agents so that electron transfer occurs thru an external wire.
- What is needed – Spontaneous Reaction
  - Salt Bridge
    - connects the 2 containers & provides a passage for ions from one solution to another
  - Electrodes
    - one of 2 surfaces that conduct electricity, the site of oxidation & reduction rxns





Flow of Electrons



Reduction occurs at the cathode

Oxidation occurs at the anode

From Anode to Cathode

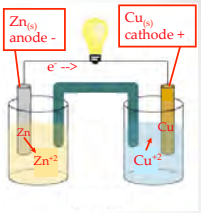
Voltaic Cell

Overall Redox Rxn

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

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

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



Species oxidized Zn

Reducing agent Zn

Species reduced Cu^{+2}

Oxidizing agent Cu^{+2}

Anode Zn

Cathode Cu

Sketch Notes

Name: _____ Date: _____ Per: _____

Assigning Oxidation Numbers

	Formula	Element and Oxidation Number					
1.	NO_2	**N**		**O**			
2.	KCl	**K**		**Cl**			
3.	MnO_2	**Mn**		**O**			
4.	H_2SO_4	**H**		**S**		**O**	
5.	K_3PO_4	**K**		**P**		**O**	
6.	HNO_3	**H**		**N**		**O**	
7.	Fe_2O_3	**Fe**		**O**			
8.	$CaCl_2$	**Ca**		**Cl**			
9.	$Na_2S_2O_3$	**Na**		**S**		**O**	
10.	CH_4	**C**		**H**			

Video 14.1: Oxidation Numbers

Determine the oxidation number of **EACH** element in each of the following substances:

1. $KMnO_4$
2. S_8
3. $Na_2Cr_2O_7$
4. $CuCl_2$
5. H_2SO_4
6. PO_4^{-3}
7. zinc oxide
8. water
9. calcium hydride
10. carbon dioxide

Video 14.2: Redox Reactions

Writing Half Reactions

Step 1: Balance the half reaction by atom using coefficients
Step 2: Balance the half reaction by charge using electrons (e^-)

Electrons always go on the side with the more positive oxidation #

Half rxns are written on a per atom basis. The only exception is if the half rxn contains a diatomic element

$Mg \rightarrow Mg^{+2}$

$O_2 \rightarrow O^{-2}$

$Cl^{+5} \rightarrow Cl^{-1}$

$Mn^{+2} \rightarrow Mn^{+5}$

$N^{-3} \rightarrow N_2$

$S^{+6} \rightarrow S^{-2}$

$3Cu^{+1} \rightarrow 3Cu$

$6H^{+1} \rightarrow 3H_2$

Name:

14.1 Oxidation Numbers

____ 1. What is the oxidation state of nitrogen in the compound NH_4Br?

1) –1 2) +2 3) –3 4) +4

____ 2. What is the oxidation number of sulfur in $Na_2S_2O_3$?

1) –1 2) +2 3) +6 4) +4

____ 3. Given the balanced equation representing a reaction:

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

The oxidation state of chlorine in this reaction changes from

1) –1 to +1
2) –1 to +5
3) + I to –1
4) +5 to –1

____ 4. What is the oxidation number of chromium in the chromate ion, CrO_4^{2-} ?

1) +6 2) +2 3) +3 4) +8

____ 5. What is the oxidation number assigned to manganese in $KMnO_4$?

1) +7 2) +2 3) +3 4) +4

____ 6. What is the oxidation state of nitrogen in $NaNO_2$?

1) +1 2) +2 3) +3 4) +4

____ 7. What is the oxidation number of chromium in $K_2Cr_2O_7$?

1) +12 2) +2 3) +3 4) +6

____ 8. Given the reaction that occurs in an electrochemical cell:

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

During this reaction, the oxidation number of Zn changes from

1) 0 to +2
2) 0 to –2
3) +2 to 0
4) –2 to 0

____ 9. In which substance does hydrogen have an oxidation number of zero?

1) LiH
2) H_2O
3) H_2S
4) H_2

____ 10. In which compound does carbon have an oxidation state of –4?

1) CO
2) CO_2
3) CCl_4
4) CH_4

Creating Half Reactions from Redox Reactions

Step 1: Assign oxidation numbers to determine the substances that have a change in oxidation number
Step 2: Identify the two half reactions & balance each half reaction
Step 3: Label each half reaction & identify the species oxidized, the species reduced, the oxidizing agent & the reducing agent

Note: The reaction provided to you may be balanced or unbalanced. Remember that half reactions are made on a per atom basis. The ONLY exception is a half reaction that contains a diatomic element

$$CuSO_4 + Mn \rightarrow Mn(SO_4)_2 + Cu$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$AgF + Fe \rightarrow FeF_3 + Ag$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$Na + Cl_2 \rightarrow NaCl$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$H_2 + O_2 \rightarrow H_2O$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$S + O_2 \rightarrow SO_3$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$Fe + HCl \rightarrow H_2 + FeCl_3$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

$$KClO_3 \rightarrow KCl + O_2$$

Species oxidized _____ Oxidizing Agent _____

Species reduced _____ Reducing Agent _____

Oxidation: Reduction:

Name:

14.2 Half Reactions

1. An oxidation-reduction reaction involves the

1) sharing of electrons
2) sharing of protons
3) transfer of electrons
4) transfer of protons

2. During which process does an atom gain one or more electrons?

1) transmutation
2) reduction
3) oxidation
4) neutralization

3. Which half-reaction correctly represents reduction?

1) $Mn^{4+} \rightarrow Mn^{3+} + e^-$
2) $Mn^{4+} \rightarrow Mn^{7+} + 3e^-$
3) $Mn^{4+} + e^- \rightarrow Mn^{3+}$
4) $Mn^{4+} + 3e^- \rightarrow Mn^{7+}$

4. In a redox reaction, the total number of electrons lost is

1) less than the total number of electrons gained
2) greater than the total number of electrons gained
3) equal to the total number of electrons gained
4) equal to the total number of protons gained

5. Which ion is most easily reduced?

1) Zn^{2+} 2) Mg^{2+} 3) Co^{2+} 4) Ca^{2+}

6. Given the balanced ionic equation representing a reaction:

$2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$

Which half-reaction represents the reduction that occurs?

1) $Al \rightarrow Al^{3+} + 3e^-$
2) $Al^{3+} + 3e^- \rightarrow Al$
3) $Cu \rightarrow Cu^{2+} + 2e^-$
4) $Cu^{2+} + 2e^- \rightarrow Cu$

7. Which equation represents an oxidation-reduction reaction?

1) $H^+ + OH^- \rightarrow H_2O$
2) $^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$
3) $Zn + Sn^{4+} \rightarrow Zn^{2+} + Sn^{2+}$
4) $3AgNO_3 + Li_3PO_4 \rightarrow Ag_3PO_4 + 3LiNO_3$

8. Which half-reaction shows conservation of charge?

1) $Cu + e^- \rightarrow Cu^+$
2) $Cu^{2+} + 2e^- \rightarrow Cu$
3) $Cu^+ \rightarrow Cu + e^-$
4) $Cu^{2+} \rightarrow Cu + 2e^-$

9. Which balanced equation represents an oxidation-reduction reaction?

1) $Ba(NO_3)_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaNO_3$
2) $H_3PO_4 + 3KOH \rightarrow K_3PO_4 + 3H_2O$
3) $Fe(s) + S(s) \rightarrow FeS(s)$
4) $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

10. Which expression correctly represents a balanced reduction half-reaction?

1) $Na^+ + e^- \rightarrow Na$
2) $Na \rightarrow Na^+ + e^-$
3) $Cl_2 + 2e^- \rightarrow Cl^-$
4) $2\,Cl^- \rightarrow Cl_2 + 2e^-$

Lesson 14.3: Activity Series and Spontaneity

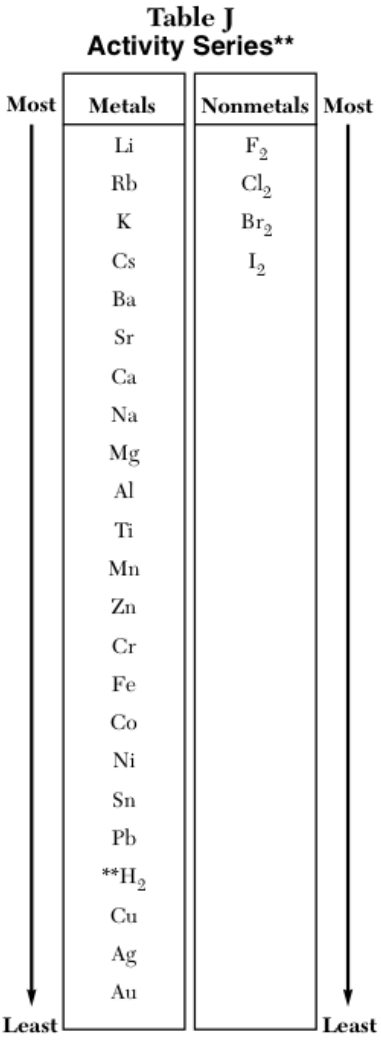


Table J: The table is arranged with the most reactive metals at the top of the table and the least reactive metals at the bottom of the table. What this means is that a metal listed on the table will react with the compound of a metal that is below it. For example, Zn is above Cu on the table. This means that Zn will replace Cu in a compound. Zn will be oxidized by a compound containing Cu. The Cu^{+} ion will be reduced by elemental Zn^{0}. Metals higher on table J are more likely to be oxidized.

The reaction: $Zn + Cu(NO_3)_2 \rightarrow Cu + Zn(NO_3)_2$ will spontaneously occur.

For each example below, if a reaction will occur spontaneously based on the element positions in the Activity Series, complete the equation and balance it. If there is no reaction, write no reaction.

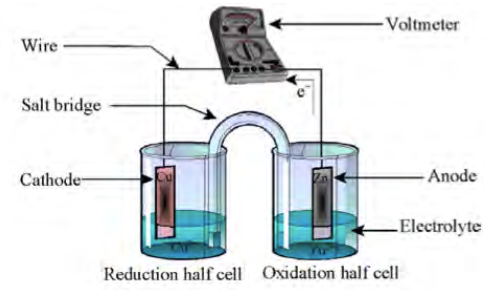
1. $Ag_{(s)} + Cu(NO_3)_{2(aq)} \rightarrow$
2. $Pb_{(s)} + AuCl_{(aq)} \rightarrow$
3. $Au_{(s)} + LiCl_{(aq)} \rightarrow$
4. $Mg_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow$

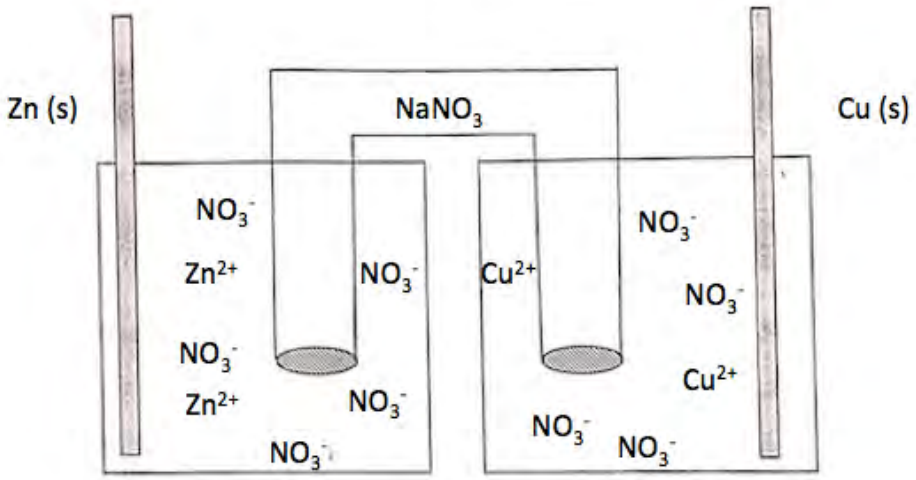
Determine if a spontaneous reaction will occur. If one will occur, write the balanced oxidation and reduction half reactions.

1. $Zn + Pb(NO_3)_2 \rightarrow Pb + Zn(NO_3)_2$
   - Is the reaction spontaneous? _____
   - If yes, write the oxidation ½ reaction:
   - If yes, write the reduction ½ reaction:
2. $3\ Zn + 2\ Al(NO_3)_3 \rightarrow 2\ Al + 3\ Zn(NO_3)_2$
   - Is the reaction spontaneous? _____
   - If yes, write the oxidation ½ reaction:
   - If yes, write the reduction ½ reaction:

Video 14.4: Electrochemical Cells

A voltaic cell is an electrochemical cell. A voltaic cell produces electric current as a result of a spontaneous redox reaction, used to make batteries. They consist of two half-cells connected by a salt bridge and two electrodes that connect to a load that uses the electricity produced by the cell.) An example of an electrochemical cell is a battery.





Overall Redox Reaction:

Oxidation ½ Reaction:

Species Oxidized: _____

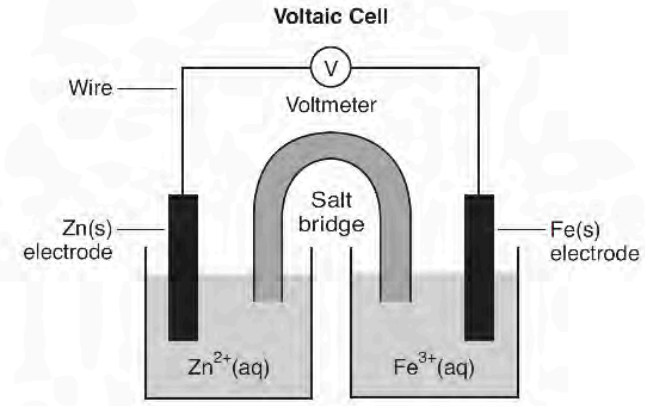
Species Reduced: _____

Reduction ½ Reaction:

Anode: _____

Cathode: _____

An operating voltaic cell has zinc and iron electrodes. The cell and the unbalanced ionic equation representing the reaction that occurs in the cell are shown below.



$$Zn(s) + Fe^{3+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$$

1. Write the balanced oxidation half-reaction for this cell.

2. Write the balanced reduction half reaction for this cell.

3. If 6 moles of Zn react, how many moles of electrons will be transferred?

4. Explain, in terms of Zn atoms and Zn ions, why the mass of the Zn electrode decreases as the cell operates.

5. Identify the subatomic particles that flow through the wire as the cell operates.

LESSON 14.4: Electrochemical Cells

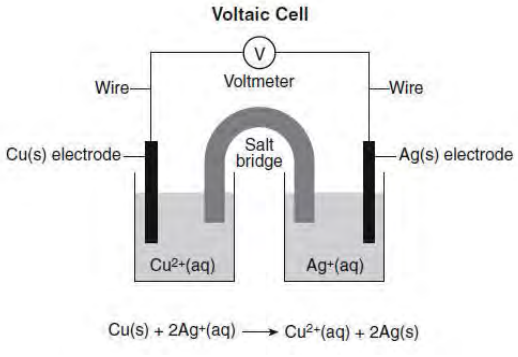
Objective:

- ***Determine the flow of electrons in a battery (voltaic cell)***
- ***Identify the anode and cathode in a voltaic cell***

Voltaic Cells (batteries)

<u>Directions</u>: In each of the following, determine which element oxidized easier on table J (higher up on table J). Then label the ***anode, cathode, direction of e- flow,*** *(remember electrons flow from high to low),* ***which electrode increases and decreases in mass*** *and then* ***write the half reactions in the spaces provided.***

1.

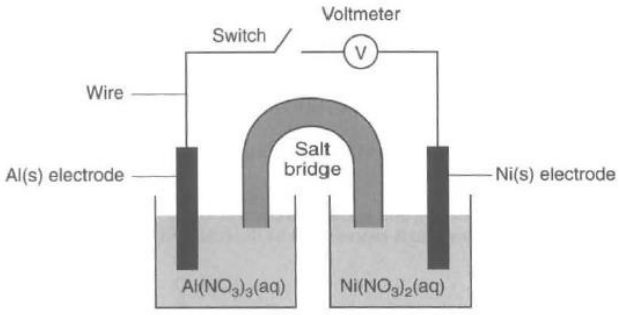


$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$

ox:

red:

2.

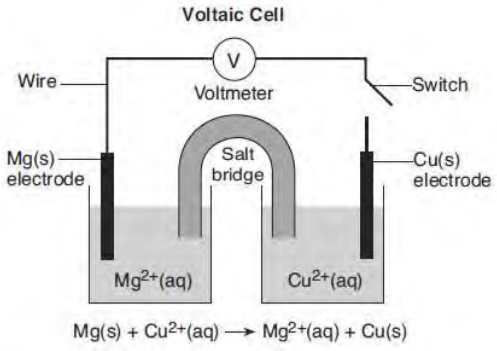


$2Al(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni(s)$

ox:

red:

3.

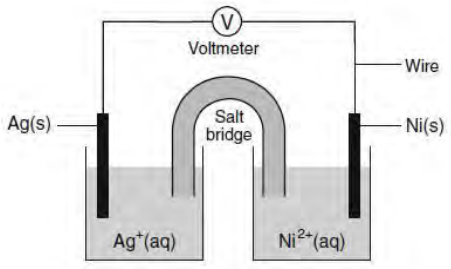


$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$

ox:

red:

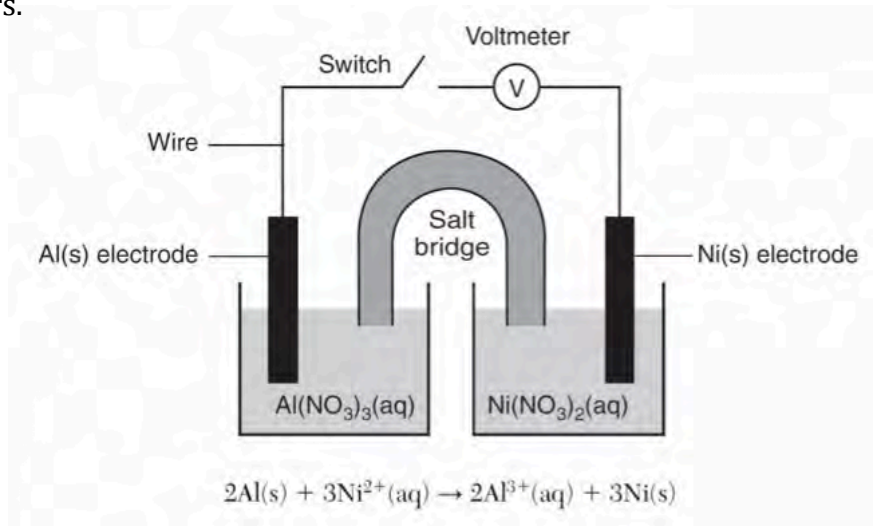
4.



ox:

red:

A student constructs an electrochemical cell during a laboratory investigation. When the switch is closed, electrons flow through the external circuit. The diagram below represents this cell and the reaction that occurs.



$$2Al(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni(s)$$

1. State, in terms of energy, why this cell is a voltaic cell.

   _____

   _____

2. Determine the number of moles of $Al_{l(s)}$ needed to completely react with 9.0 moles of $Ni^{2+}{}_{(aq)}$ ions.

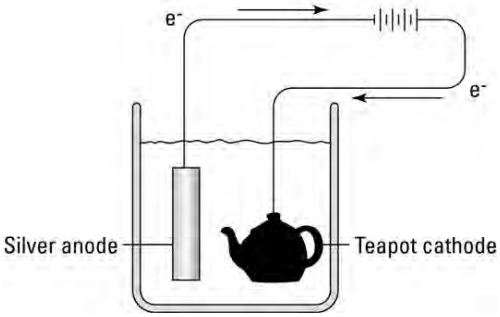
3. Write the balanced half-reaction equation for the oxidation that occurs when the switch is closed.

4. State the direction of electron flow through the wire and when the switch is closed.

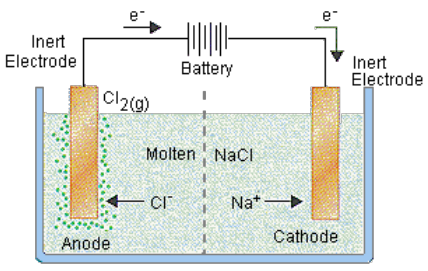
   _____

Lesson 14.5: Electrolytic Cells.

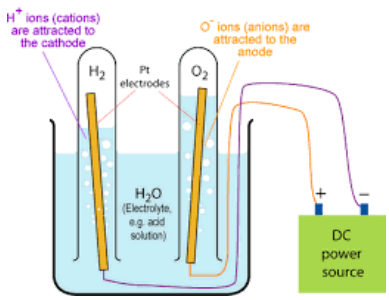
Electrolytic cells use electrical energy to force a nonspontaneous chemical reaction to occur. Reduction occurs at the cathode. The ***cathode*** is the electrode where electrons are sent and is the ***negative electrode***. Oxidation occurs at the anode. The ***anode*** is the electrode where electrons are drawn away from and is the ***positive electrode.***



Electroplating

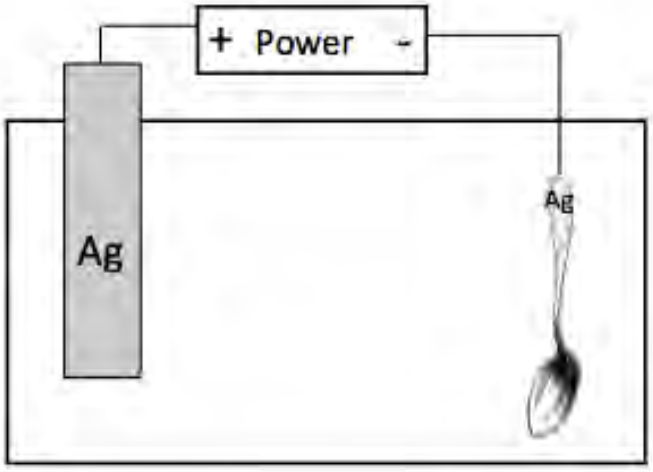


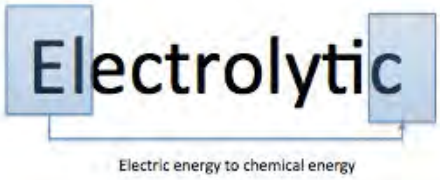
Molten Salt



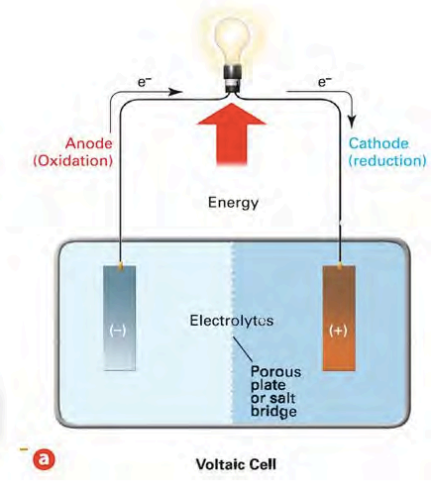
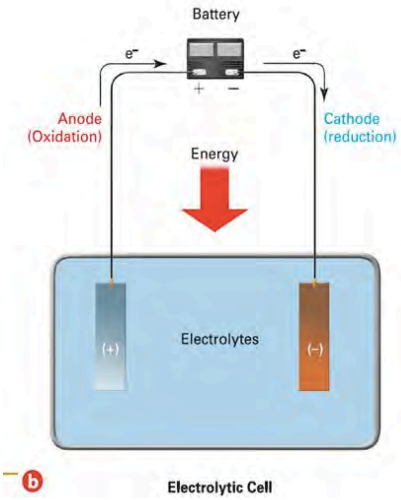
Electrolysis of Water

Electroplating:



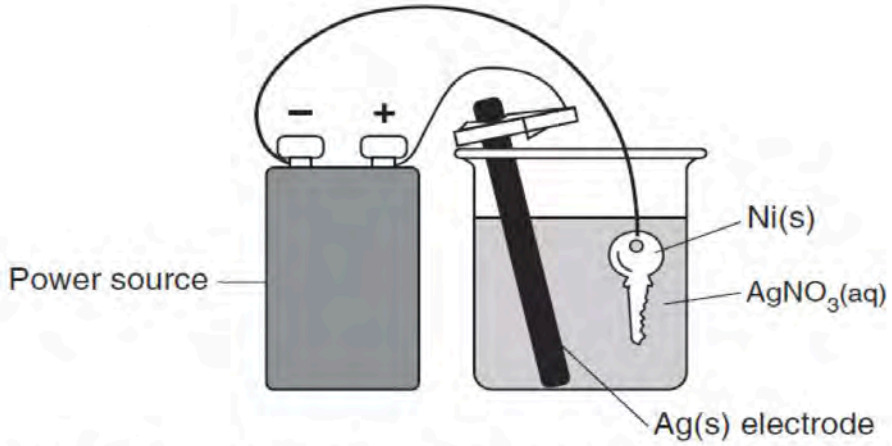


Complete the chart

Electrochemical Cells	
Voltaic Cell	***Electrolytic Cell***
Example:	Example: Electroplating
Redox Reactions	
	2 electrodes
Anode = oxidation Negative (-)	Anode =
Cathode =	Cathode = Reduction Negative (-)
Electrons flow from anode to cathode	
	Use electrical energy for force a nonspontaneous redox reaction to occur (chemical)
	

Base your answers to questions **1** through **3** on the information below.

The diagram below represents an operating electrolytic cell used to plate silver onto a nickel key. As the cell operates, oxidation occurs at the silver electrode and the mass of the silver electrode decreases.



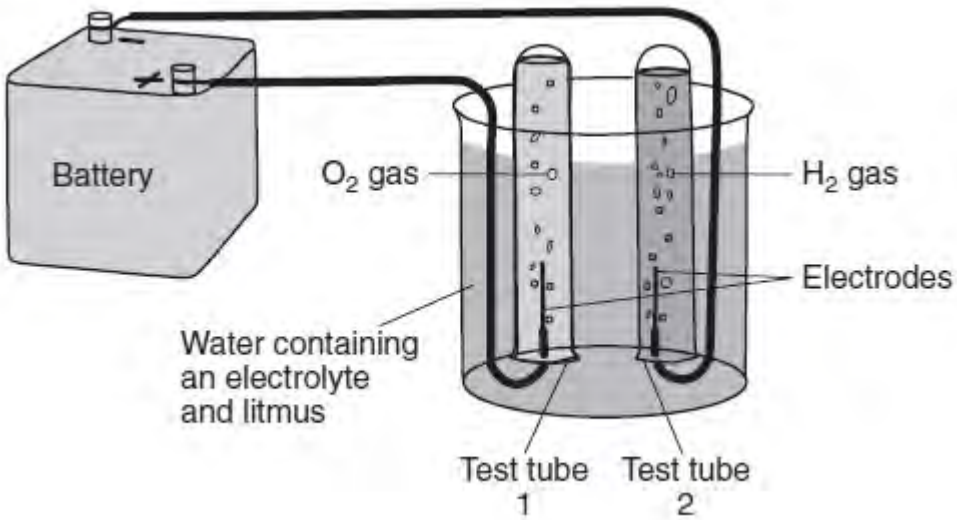
1. Explain, in terms of Ag atoms and $Ag^+(aq)$ ions, why the mass of the silver electrode *decreases* as the cell operates.

2. State the purpose of the power source in the cell.

3. Identify the cathode in the cell.

Base your answers to questions **4** and **5** on the information below.

The diagram below shows a system in which water is being decomposed into oxygen gas and hydrogen gas. Litmus is used as an indicator in the water. The litmus turns red in test tube 1 and blue in test tube 2.



The oxidation and reduction occurring in the test tubes are represented by the balanced equations below.

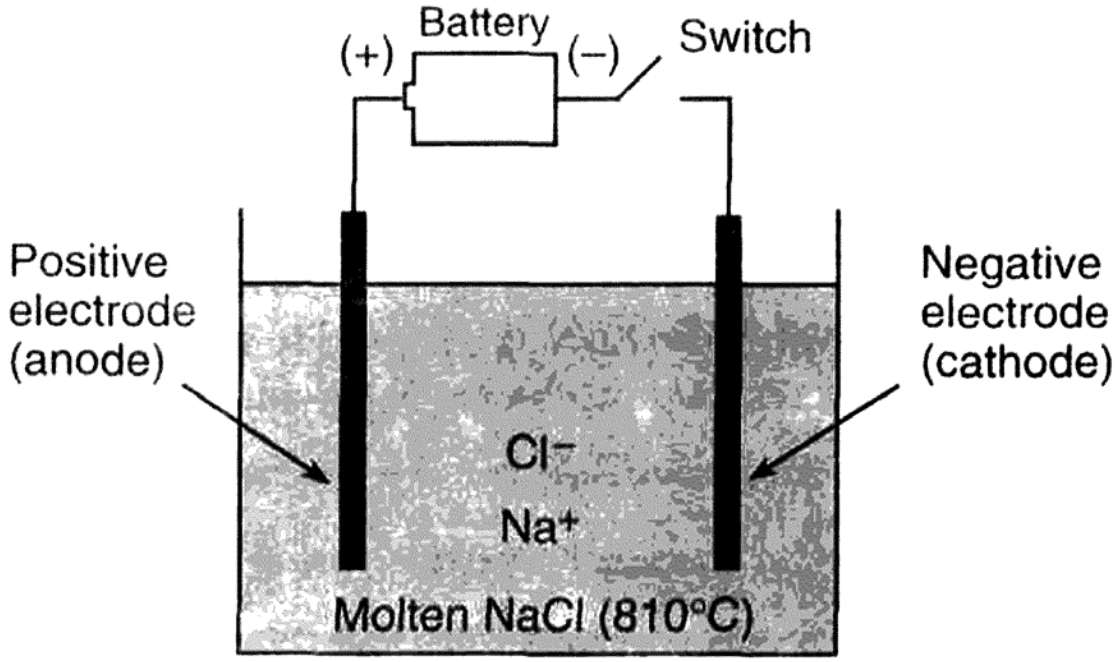
Test tube 1: $2H_2O(\ ) \quad O_2(g) + 4H^+(aq) + 4e^-$

Test tube 2: $4H_2O(\ ) + 4e^- \quad 2H_2(g) + 4OH^-(aq)$

4. Determine the change in oxidation number of oxygen during the reaction in test tube 1.

5. Identify the information in the diagram that indicates this system is an electrolytic cell.

Base your answers to questions **6** through **8** on the diagram and balanced equation below, which represent the electrolysis of molten NaCl.



$$2NaCl \longrightarrow Cl_2 + 2Na$$

6. What is the purpose of the battery in this electrolytic cell?

7. When the switch is closed, which electrode will attract the sodium ions?

8. Write the balanced half-reaction for the reduction that occurs in this electrolytic cell.

Redox Review

1. What is the oxidation state of nitrogen in the compound NH_4Br?

   A) –1 B) +2 C) –3 D) +4

2. Given the balanced equation representing a reaction:

   $$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

   The oxidation state of chlorine in this reaction changes from

   A) –1 to +1 B) –1 to +5
   C) + l to –1 D) +5 to –1

3. During which process does an atom gain one or more electrons?

   A) transmutation B) reduction
   C) oxidation D) neutralization

4. Which half-reaction correctly represents reduction?

   A) $Mn^{4+} \rightarrow Mn^{3+} + e^-$
   B) $Mn^{4+} \rightarrow Mn^{7+} + 3e^-$
   C) $Mn^{4+} + e^- \rightarrow Mn^{3+}$
   D) $Mn^{4+} + 3e^- \rightarrow Mn^{7+}$

5. In a redox reaction, the total number of electrons lost is

   A) less than the total number of electrons gained
   B) greater than the total number of electrons gained
   C) equal to the total number of electrons gained
   D) equal to the total number of protons gained

6. Half-reactions can be written to represent all

   A) double-replacement reactions
   B) neutralization reactions
   C) fission and fusion reactions
   D) oxidation and reduction reactions

7. Which half-reaction equation represents the reduction of an iron(II) ion?

   A) $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   B) $Fe^{2+} + 2e^- \rightarrow Fe$
   C) $Fe^{3+} + e^- \rightarrow Fe^{2+}$
   D) $Fe \rightarrow Fe^{2+} + 2e^-$

8. Given the balanced equation representing a reaction:

   $$Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$$

   During this reaction, the oxidation number of Fe changes from

   A) +2 to 0 as electrons are transferred
   B) +2 to 0 as protons are transferred
   C) +3 to 0 as electrons are transferred
   D) +3 to 0 as protons are transferred

9. Given the balanced equation representing a redox reaction:

   $$2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$$

   Which statement is true about this reaction?

   A) Each Al loses $2e^-$ and each Cu^{2+} gains $3e^-$.
   B) Each Al loses $3e^-$ and each Cu^{2+} gains $2e^-$.
   C) Each Al^{3+} gains $2e^-$ and each Cu loses $3e^-$.
   D) Each Al^{3+} gains $3e^-$ and each Cu loses $2e^-$.

10. Which half-reaction correctly represents reduction?

    A) $Ag \rightarrow Ag^+ + e^-$
    B) $F_2 \rightarrow 2\ F^- + 2e^-$
    C) $Au^{3+} + 3e^- \rightarrow Au$
    D) $Fe^{2+} + e^- \rightarrow Fe^{3+}$

11. Which balanced equation represents a redox reaction?

    A) $AgNO_3(aq) + NaCI(aq) \rightarrow AgCI(s) + NaNO_3(aq)$
    B) $H_2CO_3(aq) \rightarrow H_2O(\ell) + CO_2(g)$
    C) $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(\ell)$
    D) $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

12. Which equation represents an oxidation-reduction reaction?

    A) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
    B) $H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$
    C) $MgCrO_4 + BaCl_2$ ® $MgC1_2 + BaCrO_4$
    D) $Zn(NO_3)_2 + Na_2CO_3 \rightarrow 2NaNO_3 + ZnCO_3$

13. Which metal is more active than H_2?

A) Ag B) Au C) Cu D) Pb

14. Which reaction occurs spontaneously?

A) $Cl_2(g) + 2NaBr(aq) \rightarrow Br_2(\ell) + 2NaCl(aq)$
B) $Cl_2(g) + 2NaF(aq) \rightarrow F_2(g) + 2NaCl(aq)$
C) $I_2(s) + 2NaBr(aq) \rightarrow Br_2(\ell) + 2NaI(aq)$
D) $I_2(s) + 2NaF(aq) \rightarrow F_2(g) + 2NaI(aq)$

15. Given the unbalanced ionic equation:

$$3Mg + \_\_Fe^{3+} \rightarrow 3Mg^{2+} + \_\_Fe$$

When this equation is balanced, both Fe^{3+} and Fe have a coefficient of

A) 1, because a total of 6 electrons is transferred
B) 2, because a total of 6 electrons is transferred
C) 1, because a total of 3 electrons is transferred
D) 2, because a total of 3 electrons is transferred

16. Which half-reaction shows conservation of charge?

A) $Cu + e^- \rightarrow Cu^+$
B) $Cu^{2+} + 2e^- \rightarrow Cu$
C) $Cu^+ \rightarrow Cu + e^-$
D) $Cu^{2+} \rightarrow Cu + 2e^-$

17. Given the balanced equation representing a reaction:

$$2Fe + 3Cu^{2+} \rightarrow 2Fe^{3+} + 3Cu$$

When the iron atoms lose six moles of electrons, how many moles of electrons are gained by the copper ions?

A) 12 moles B) 2 moles
C) 3 moles D) 6 moles

18. A voltaic cell spontaneously converts chemical energy to

A) electrical energy
B) geothermal energy
C) mechanical energy
D) nuclear energy

19. In a voltaic cell, chemical energy is converted to

A) electrical energy, spontaneously
B) electrical energy, non-spontaneously
C) nuclear energy, spontaneously
D) nuclear energy, non-spontaneously

20. Which half-reaction can occur at the anode in a voltaic cell?

A) $Ni^{2+} + 2e^- \rightarrow Ni$
B) $Sn + 2e^- \rightarrow Sn^{2+}$
C) $Zn \rightarrow Zn^{2+} + 2e^-$
D) $Fe^{3+} \rightarrow Fe^{2+} + e^-$

21. Given the balanced ionic equation representing the reaction in an operating voltaic cell:

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

The flow of electrons through the external circuit in this cell is from the

A) Cu anode to the Zn cathode
B) Cu cathode to the Zn anode
C) Zn anode to the Cu cathode
D) Zn cathode to the Cu anode

22. Reduction occurs at the cathode in

A) electrolytic cells, only
B) voltaic cells, only
C) both electrolytic cells and voltaic cells
D) neither electrolytic cells nor voltaic cells

23. Given the balanced equation representing the reaction occurring in a voltaic cell:

$$Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$$

In the completed external circuit, the electrons flow from

A) Pb(s) to Zn(s)
B) $Pb^{2+}(aq)$ to $Zn^{2+}(aq)$
C) Zn(s) to Pb(s)
D) $Zn^{2+}(aq)$ to $Pb^{2+}(aq)$

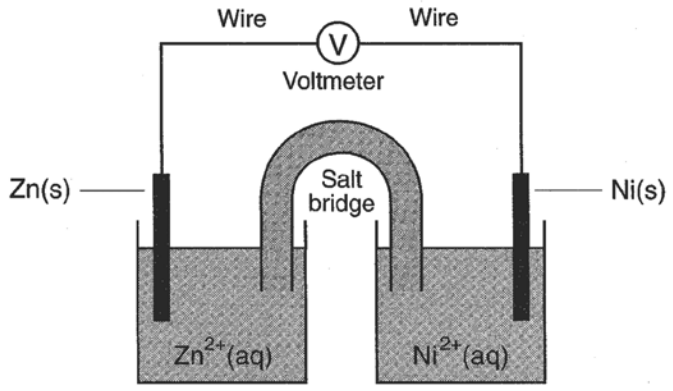
24. Which statement is true about oxidation and reduction in an electrochemical cell?

A) Both occur at the anode.
B) Both occur at the cathode.
C) Oxidation occurs at the anode and reduction occurs at the cathode.
D) Oxidation occurs at the cathode and reduction occurs at the anode.

25. Which statement describes one characteristic of an operating electrolytic cell?

A) It produces electrical energy.
B) It requires an external energy source.
C) It uses radioactive nuclides.
D) It undergoes a spontaneous redox reaction.

26. The diagram below represents an operating electrochemical cell and the balanced ionic equation for the reaction occurring in the cell.



$Zn(s) + Ni^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Ni(s)$

Which statement identifies the part of the cell that conducts electrons and describes the direction of electron flow as the cell operates?

A) Electrons flow through the salt bridge from the Ni(s) to the Zn(s).
B) Electrons flow through the salt bridge from the Zn(s) to the Ni(s).
C) Electrons flow through the wire from the Ni(s) to the Zn(s).
D) Electrons flow through the wire from the Zn(s) to the Ni(s).

27. Which statement describes electrolysis?

A) Chemical energy is used to produce an electrical change.
B) Chemical energy is used to produce a thermal change.
C) Electrical energy is used to produce a chemical change.
D) Thermal energy is used to produce a chemical change.

28. Given the balanced equation representing a reaction occurring in an electrolytic cell:

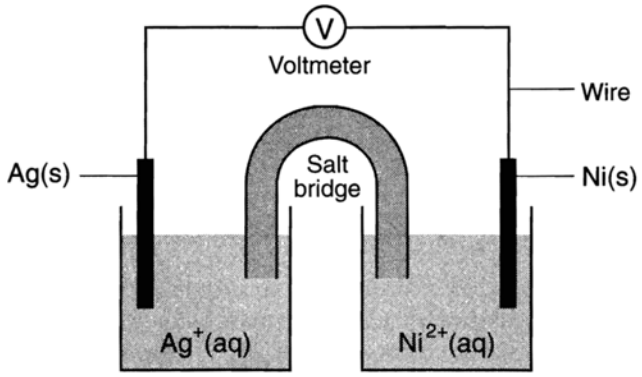
$$2NaCl(\ell) \rightarrow 2Na(\ell) + Cl_2(g)$$

Where is $Na(\ell)$ produced in the cell?

A) at the anode, where oxidation occurs
B) at the anode, where reduction occurs
C) at the cathode, where oxidation occurs
D) at the cathode, where reduction occurs

Base your answers to questions **29** through **31** on the information below.

The diagram below represents an operating voltaic cell at 298 K and 1.0 atmosphere in a laboratory investigation. The reaction occurring in the cell is represented by the balanced ionic equation below.



$$2Ag^+(aq) + Ni(s) \longrightarrow 2Ag(s) + Ni^{2+}(aq)$$

29. Determine the total number of moles of $Ni^{2+}(aq)$ ions produced when 4.0 moles of $Ag^+(aq)$ ions completely react in this cell

30. Identify the anode in this cell.

31. Write a balanced half-reaction equation for the reduction that occurs in this cell.

32. Base your answer to the following question on the information below.

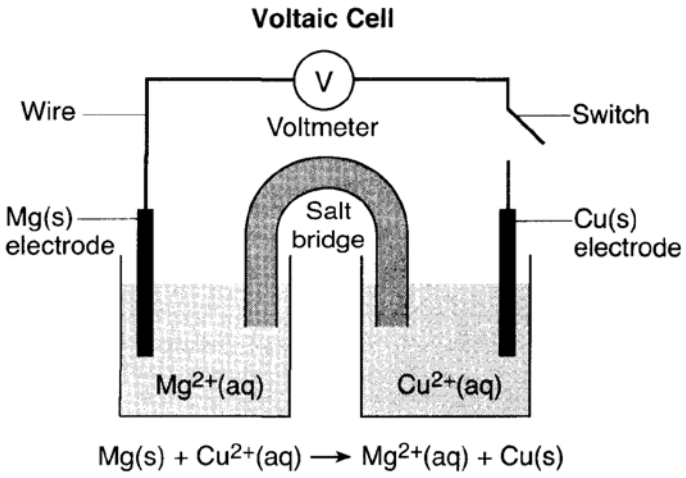
In a laboratory investigation, magnesium reacts with hydrochloric acid to produce hydrogen gas and magnesium chloride. This reaction is represented by the unbalanced equation below.

$$Mg(s) + HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$$

Write a balanced half-reaction equation for the oxidation that occurs.

Base your answers to questions **33** through **35** on the information below.

A voltaic cell with magnesium and copper electrodes is shown in the diagram below. The copper electrode has a mass of 15.0 grams.



When the switch is closed, the reaction in the cell begins. The balanced ionic equation for the reaction in the cell is shown below the cell diagram. After several hours, the copper electrode is removed, rinsed with water, and dried. At this time, the mass of the copper electrode is greater than 15.0 grams.

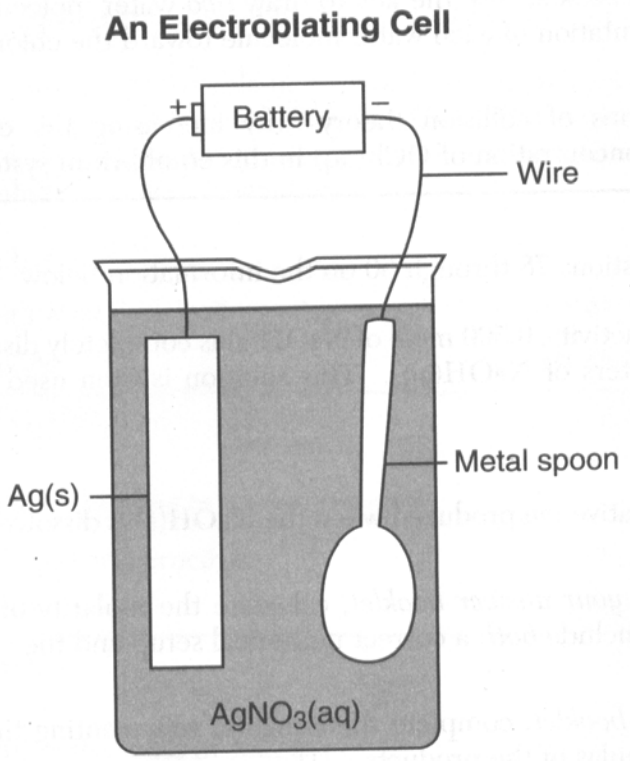
33. State the purpose of the salt bridge in this cell.

34. State the directions of electron flow through the wire between the electrodes when the switch is closed.

35. Explain, in terms of copper ions and copper atoms, why the mass of the copper electrode increases as the cell operates. Your response must include information about *both* copper ions and copper atoms.

36. Base your answer to the following question on the information below.

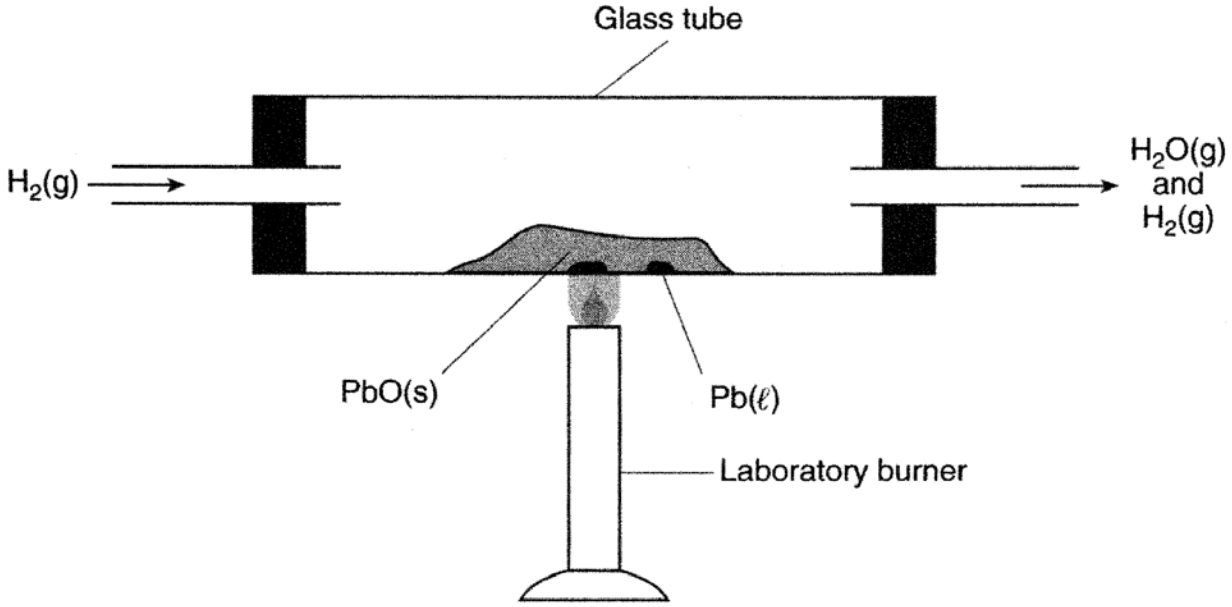
Electroplating is an electrolytic process used to coat metal objects with a more expensive and less reactive metal. The diagram below shows an electroplating cell that includes a battery connected to a silver bar and a metal spoon. The bar and spoon are submerged in $AgNO_3(aq)$.



Explain the purpose of the battery in this cell.

Base your answers to questions **37** through **40** on the information below and on your knowledge of chemistry.

In a laboratory apparatus, a sample of lead(II) oxide reacts with hydrogen gas at high temperature. The products of t his reaction are liquid lead and water vapor. As the reaction proceeds, water vapor and excess hydrogen gas leave the glass tube. The diagram and balanced equation below represent this reaction.



$$PbO(s) + H_2(g) + \text{heat} \longrightarrow Pb(\ell) + H_2O(g)$$

37. State *one* change in reaction conditions, other than adding a catalyst, that would cause the rate of this reaction to increase.

38. Explain why the reaction that occurs in this glass tube can *not* reach equilibrium.

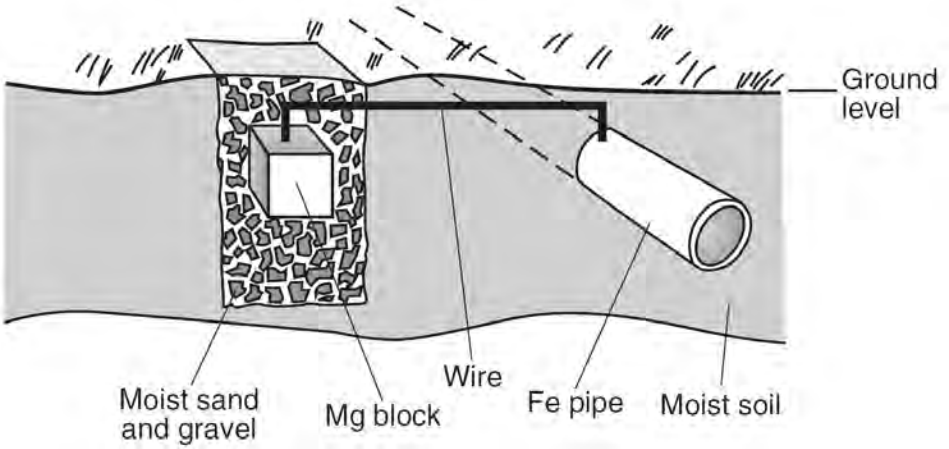
39. Write a balanced half-reaction equation for the reduction of the Pb^{2+} ions in this reaction.

40. Determine the change in oxidation number for the hydrogen that reacts.

Base your answers to questions **41** and **42** on the information below.

Underground iron pipes in contact with moist soil are likely to corrode. This corrosion can be prevented by applying the principles of electrochemistry. Connecting an iron pipe to a magnesium block with a wire creates an electrochemical cell. The magnesium block acts as the anode and the iron pipe acts as the cathode. A diagram of this system is shown below.

Cross-Sectional View of Underground Pipe Protection System

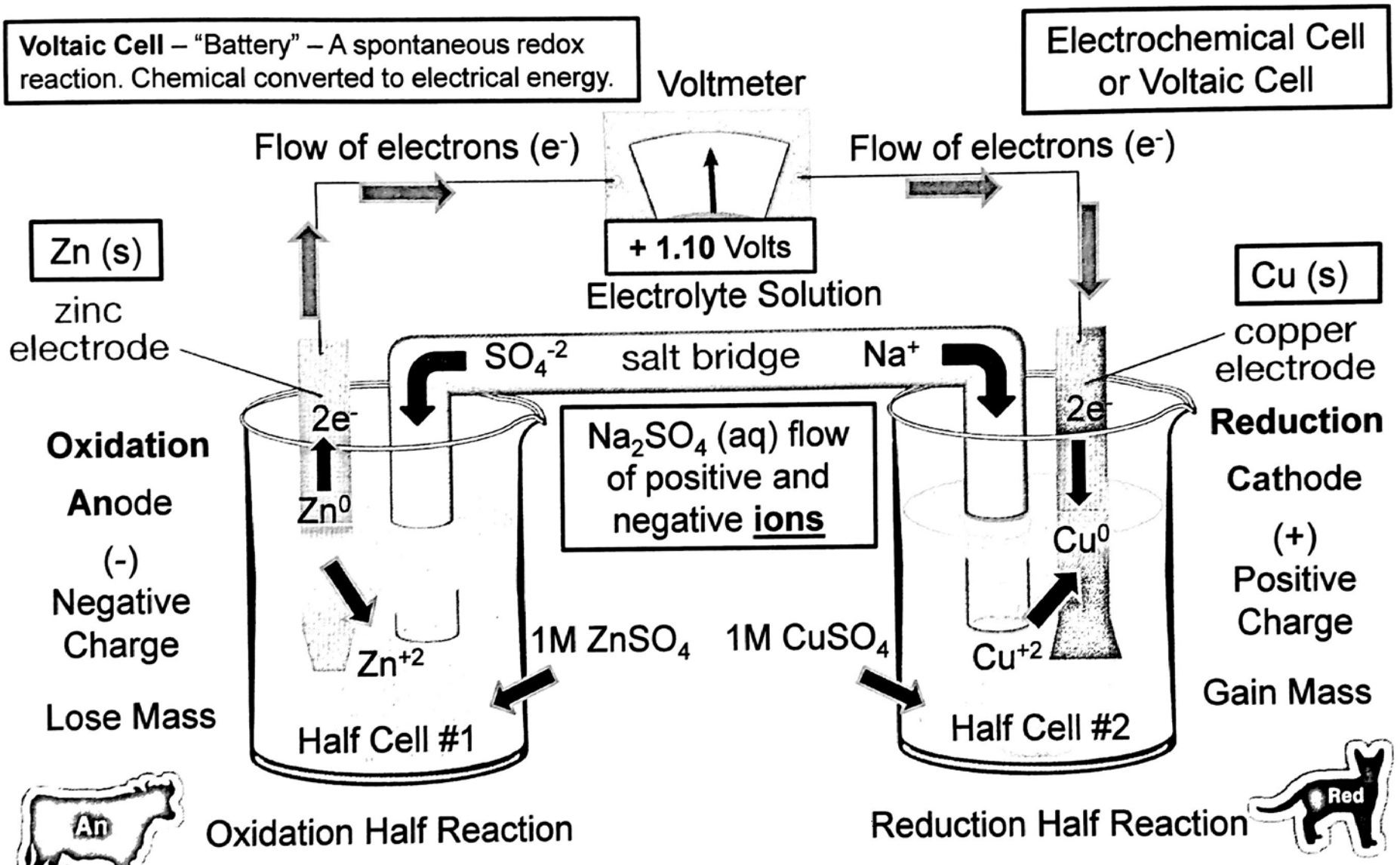


41. Explain, in terms of reactivity, why magnesium is preferred over zinc to protect underground iron pipes. Your response must include *both* magnesium and zinc.

42. State the direction of the flow of electrons between the electrodes in this cell.

Voltaic Cell – "Battery" – A spontaneous redox reaction. Chemical converted to electrical energy.

Electrochemical Cell or Voltaic Cell



Oxidation Half Reaction

Zn^0 (s) → Zn^{+2} (aq) + $2e^-$

E^0 = + 0.76

Reduction Half Reaction

Cu^{+2} (aq) + $2e^-$ → Cu^0 (s)

E^0 = + 0.34

Net ionic reaction: Zn^0 (s) + Cu^{+2} (aq) → Zn^{+2} (aq) + Cu^0 (s)