REDOX

BASICS + Notes

Arbuiso Chem

Read all of this. Know all of this.

Name: _____

REDOX BASICS - Oxidation & Reduction Chemistry

<u>According to NY State Regents Chem Guidelines, this is what we have to learn</u>

1. An oxidation-reduction (redox) reaction involves the transfer of electrons (e^-)

2. Oxidation is the loss of electrons (<u>LEO</u> the Lion goes GER).

3. A half-reaction can be written to represent oxidation.

4. Reduction is the gain of electrons (LEO the Lion goes <u>GER</u>).

5. A half-reaction can be written to represent reduction.

6.   In a redox reaction, the number of electrons lost = the number of electrons gained.

7.   Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers
       indicate that oxidation and reduction have occurred.

8. An electrochemical cell can either be voltaic or electrolytic. In an electrochemical cell, oxidation
     occurs at the anode and reduction at the cathode (Leo's a RED-CAT)

9.   A voltaic cell spontaneously converts chemical energy to electrical energy.

 10. An electrolytic cell requires electrical energy to produce chemical change. This process is known as
       electrolysis.

1. An oxidation-reduction (redox) reaction involves the transfer of electrons

We've learned already of many kinds of reactions in chemistry; Synthesis, decomposition, single replacement, double replacement, & combustion. Then acid-base neutralization. Finally in organic chem we learned about addition and substitution, esterification, polymerization, fermentation & saponification.

That's 12 so far (and several more await in nuclear chem).

REDOX is the next reaction, but it's not really a separate one, many of the reactions above are ALSO redox reaction.
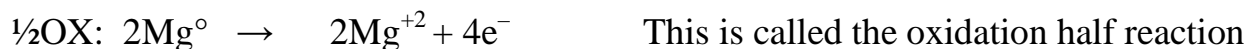
Redox reactions occur whenever there is a transfer of electrons, when atoms become cations or anions, or the reverse, when ions become atoms.

2. Oxidation is the loss of electrons (LEO the Lion goes GER).
3. A half-reaction can be written to represent oxidation.

$$2Mg + O_2 \ \rightarrow \ 2MgO$$
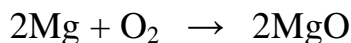
This is a common reaction for us, a synthesis that we did in lab earlier in the year.

What's going on with the electrons here? Let's take a close look.

$2Mg \ + O_2 \ \rightarrow \ 2MgO$ Looking only at the Mg now...

½OX: $2Mg° \ \rightarrow \ 2Mg^{+2} + 4e^-$ This is called the oxidation half reaction

   Both of the magnesium atoms become +2 ions by losing 2 electrons each. (that's 4 electrons total)

4. Reduction is the gain of electrons (LEO the Lion goes <u>GER</u>).
5. A half-reaction can be written to represent reduction.

$2Mg \ + O_2 \ \rightarrow \ 2MgO$ Looking only at the oxygen now...

½RED: $O_2° \ + \ 4e^- \ \rightarrow \ 2O^{-2}$ This is called the oxidation half reaction

At the same time the magnesium atoms become +2 cations, the oxygen atoms become –2 anions.

By combining these two reactions, we have an oxidation reaction and a reduction reaction, that are perfectly balanced. The number of electrons that are oxidized must also be reduced.
There are no left over electrons, or IOU electrons.

There's just one easy rule to follow:

Make sure that you balance your oxidation & reductions.

For every single electron that is oxidized off, it has to be picked up by some other atom or ion and be reduced. No left over electrons ever. Not even one.

7. Oxidation numbers (states) can be assigned to atoms and ions. Changes in oxidation numbers indicate that oxidation and reduction have occurred.

Oxidation numbers were used earlier in the year when we put together various molecular compounds (remember the 5 different nitrogen/oxygen compounds, and the 2 different carbon/oxygen compounds)?

Oxidation numbers are listed on our periodic tables. Atoms always have oxidation numbers of ZERO.

Ions have oxidation numbers equal to there ionic charge. Atoms in molecular compounds can have a variety of oxidation numbers provided that all the oxidation numbers in a molecule sum to zero.
Polyatomic ions also have oxidation numbers equal to their charges. Using table E you can determine the oxidation numbers (charges) for each part of these polyatomic ions.

Selected oxidation numbers: (you MUST open your periodic table now, or just stop reading).

Our key element at the top of the page is carbon. Top right corner of the box shows 3 selected oxidation numbers. They are –4, +2, and +4. There are others, but in our class we'll only use these <u>selected</u> oxidation states on our periodic table.

All group 1 ions have a +1 oxidation state (charge). All group 2 ions have a +2 oxidation state (charge).

Transitional metals have one or more possible oxidation states (charges). That's why we need to use the roman numerals in naming some transitional metal ionic compounds.

Most of the nonmetals have many possible oxidation states, both positive or negative.

Almost all of the Noble Gases have a "0" since they do not make any compounds.

Let's look at these compounds and see how their oxidation numbers sum to zero (for compounds) or to a positive or negative charge (for the ions).

Sometimes you will meet a compound or ion that won't "work" with the selected oxidation numbers on our

compound or ion	net charge	oxidation numbers of each of the parts
NaCl	0	Na^{+1} Cl^{-1}
NaOH	0	Na^{+1} O^{-2} H^{+1}
H_2SO_4	0	H^{+1} H^{+1} S^{+6} O^{-2} O^{-2} O^{-2} O^{-2}
N_2O_5	0	N^{+5} N^{+5} O^{-2} O^{-2} O^{-2} O^{-2} O^{-2}
NH_3	0	N^{-3} H^{+1} H^{+1} H^{+1}
NH_4^{+1}	+1	N^{-3} H^{+1} H^{+1} H^{+1} H^{+1}
MnO_4^{-1}	-1	Mn^{+7} O^{-2} O^{-2} O^{-2} O^{-2}
PO_4^{-3}	-3	P^{+5} O^{-2} O^{-2} O^{-2} O^{-2}
HCO_3^{-1}	-1	H^{+1} C^{+4} O^{-2} O^{-2} O^{-2}
$Cr_2O_7^{-2}$	-2	Cr^{+6} Cr^{+6} O^{-2} O^{-2} O^{-2} O^{-2} O^{-2} O^{-2} O^{-2}

table. Remember, they are called "SELECTED" oxidation numbers, there are more of them. This is an intro class and sometimes "real" chem blurs into "regents" chemistry. If you can't make the numbers jive, ask your teacher.

Let's look at the oxidation numbers of all the species involved. Species is a biology word but the State

Education department loves it. Here, magnesium comes in 2 "species", the atom and the +2 cation.
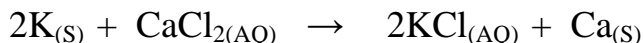Here, the oxygen is in two species as well, the atom and the oxide -2 anion.

$$2Mg + O_2 \rightarrow 2MgO \quad \text{can be thought of this way too:}$$

$$2Mg° + O_2° \rightarrow 2Mg^{+2}O^{-2}$$
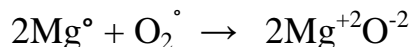
Mg atoms are $Mg°$ Oxygen molecules (a pair of atoms) are also $O_2°$
In MgO there is a Mg^{+2} cation, and the oxide anion O^{-2}

The sum of the oxidation numbers in the MgO is $(+2) + (-2) = 0$ (as expected and required)

A second example reaction:

$$2K_{(S)} + CaCl_{2(AQ)} \rightarrow 2KCl_{(AQ)} + Ca_{(S)}$$

This is a single replacement reaction, table J shows K higher than Ca, so the reaction goes forward as potassi-um has a higher activity and it will go into solution and bump out the calcium. To do this, the potassium must oxidize (or lose electrons). When this happens, the calcium ions in solution are forced to pick up these elec-trons, therefore the Ca^{+2} ions are reduced.

The redox half reactions would be:

$$\frac{1}{2}OX: \ 2K° \rightarrow 2K^{+1} + 2e^- \quad \text{note: it takes 2 K atoms to oxidize 2 electons}$$

$$\frac{1}{2}RED: \ Ca^{+2} + 2e^- \rightarrow Ca° \quad \text{note: it takes one } Ca^{+2} \text{ to reduce 2 electrons}$$

Since each half reaction is perfectly balanced we can rewrite these pair of reactions together,
omitting the electrons—since they balance out on each side of the arrow.

We can write what is called the NET IONIC EQUATION
(combining them together, after we cancel out the two electrons that are on opposite sides of the arrow).

$$2K° + Ca^{+2} \rightarrow 2K^{+1} + Ca°$$
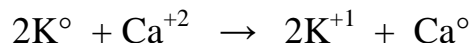
It shows both potassium atoms become K^{+1} cations, and the Ca^{+2} cation become a Ca atom.
The net ionic equation shows only the NET ion transfer inside the redox reaction. It cancels out the electrons
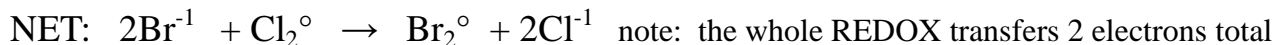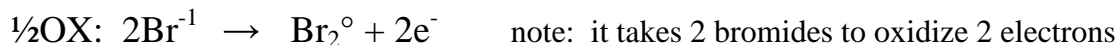from each side of the arrow.

Please take a moment to "count" the charges on both sides, they even out perfectly.

This shows the "Conservation of Charge", which is like conservation of matter and conservation
of energy. There must be conserevation of charge (or you made a mistake).

Let's look at one more reaction now…

$$Cl_2 + 2HBr \rightarrow 2HCl + Br_2$$
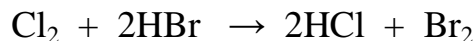
We can chemically write out both the oxidation and the reduction half reactions, and follow that with the NET IONIC EQUATION this way…

½OX: $2Br^{-1} \rightarrow Br_2{}^{\circ} + 2e^-$ note: it takes 2 bromides to oxidize 2 electrons

½RED: $Cl_2{}^{\circ} + 2e^- \rightarrow 2Cl^{-1}$ note: it takes 2 chlorine atoms to reduce 2 electrons

NET: $2Br^{-1} + Cl_2{}^{\circ} \rightarrow Br_2{}^{\circ} + 2Cl^{-1}$ note: the whole REDOX transfers 2 electrons total

In this reaction, the H^{+1} ions from the HBr, which end up HCl, are called SPECATOR IONS, because they just "watch" but don't do anything. They're required, but dull.

Single replacement reactions are also redox. Lots of reactions are redox too.

_____

Things get a little bit more involved when the number of ions oxidized by one part does not match the number gained by the other. Then, balancing reactions comes into play. For example…

$$2Al_{(S)} + 3CaCl_{2(AQ)} \rightarrow 2AlCl_{3(AQ)} + 3Ca_{(S)}$$ is a single replacement reaction
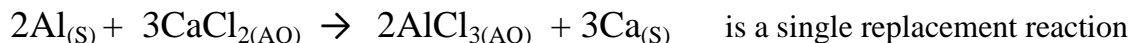
Since aluminum will oxidize here because Al is "higher" on table J compared to the Ca
(that means the Al is more reactive than Ca) It forces the calcium to become reduced.

Note: each Al loses $3e^-$ but each Ca only gains $2e^-$.
Oxidation is loss of electrons, Reduction is gain of electrons, so…

½OX: $Al \rightarrow Al^{+3} + 3e^-$
½RED: $Ca^{+2} + 2e^- \rightarrow Ca^{\circ}$
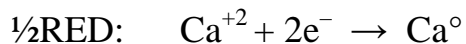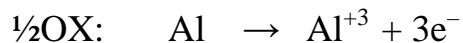
The electron transfer is NOT equal. Three electrons are oxidized, only two are used for reduction. This can't be. To "fix" this, we look for the lowest common factor (which is 6) and adjust the half reactions to match up the electrons being transferred.

½OX: $2Al \rightarrow 2Al^{+3} + 6e^-$ note: it takes 2 Al atoms to oxidize 6 electrons
½RED: $3Ca^{+2} + 6e^- \rightarrow 3Ca^{\circ}$ note: it takes 3 Ca+2 to reduce 6 electrons

Objective 8 & 9—Electrochemical Cells

An electrochemical cell can either be voltaic or electrolytic. In all electrochemical cells, oxidation occurs at the anode and reduction at the cathode (Leo's a RED-CAT)

A voltaic cell is a battery, where chemistry SPONTANEOUSLY creates electricity. These run and run (unless a switch is put into the circuit).

An electrolytic cell uses electricity to force a redox reaction that cannot happen spontaneously. These reactions are used to "electroplate" precious metals onto strong but inexpensive ones. An external battery is necessary to make this happen

This is a generic diagram for a battery, or voltaic cell. There are 2 beakers, with a metal bar in each, connected to each other by a wire. Wires carry electricity.

There is an upside down "U" tube, that contains a salt solution, connecting the two beakers.

In each beaker is an aqueous solution. Batteries that you are used to are modifications of this set up, cool technological advances in package design so that you don't have to fit two beakers into your cell phone or calculator. This is how scientists created the first batteries, the science part, not the packaging part. You're so lucky, feel it now.

BULB

NaCl aqueous

Al

Aluminum chloride
solution

Zn

Zinc Nitrate
Solution
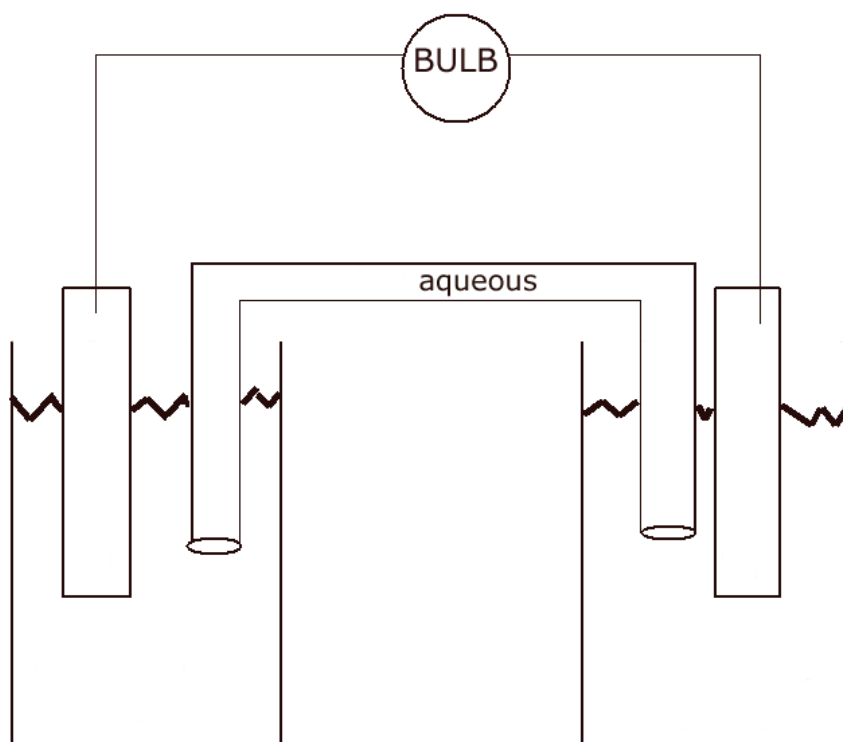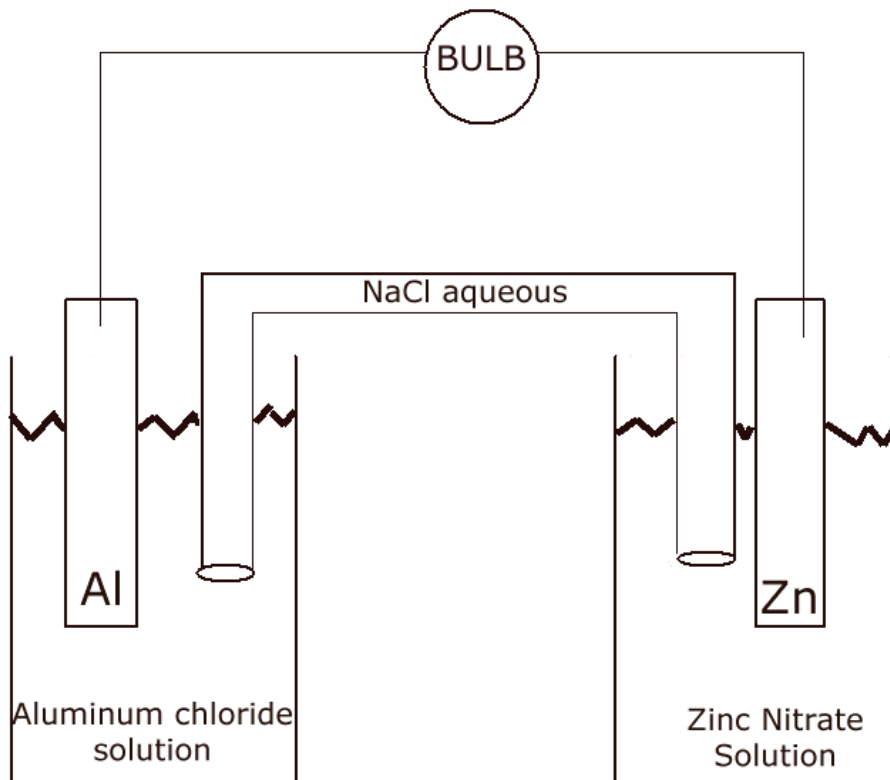
This is called an aluminum/zinc voltaic cell (battery).
Each metal bar must be in a solution which has that metal as an ion. Aluminum metal is in $AlCl_{3(AQ)}$

The Zinc is in a $Zn(NO_3)_{2(AQ)}$

There is a $NaCl_{(AQ)}$ aqueous solution in the salt bridge.

The metal bars are connected by wire, through a bulb, which lights up if electricity flows through it.

Both Al and Zn would *like to* oxidize into ions, by losing electrons. Only one will, Table J decides.

It's aluminum that will oxidize, not zinc, you know this because the Al is "higher" than Zn on table J.

When $Al \rightarrow Al^{+3}$ cations, the ions go into solution, the electrons flow up through the wire to the zinc bar.

Zinc wants to oxidize but it cannot with aluminum, so it is forced to be the reduction side of this redox.

The Zn^{+2} ions in solution are attracted to all the electrons on the zinc bar, so they jump onto the bar, gaining 2 electrons each, and are reduced into zinc atoms.

The addition of Al^{+3} cations to the left solution creates a + charge in the solution.
The removal of Zn^{+2} ions from the other solution makes a – charge in that solution.
This is a BIG PROBLEM, big enough to stop the electricity flow immediately.

That's why we have an aqueous salt bridge. In this one are Na^{+1} and Cl^{-1} ions.

As the solution at left gains cations, it becomes POSITIVELY charged, so chloride anions flow into it through the "U" tube. Those chlorides offset the positive charge, keeping that beaker neutral.

As the solution at right gets more NEGATIVELY charged because it's LOSING cations onto the zinc metal bar, the sodium cations flow to it, offsetting that charge, keeping this beaker neutral.

It's this salt bridge ion flow that keeps the beakers neutral, and allows the electrons to flow. It lets the RE-DOX to continue.

Look below and make sure you see all of this. Note the light bulb is lit because electricity (electrons) flow from the aluminum bar to the zinc bar. We have separated the two half reactions and can "use" the moving electrons to do work. Some fancy packaging and we have our own aluminum/zinc battery.
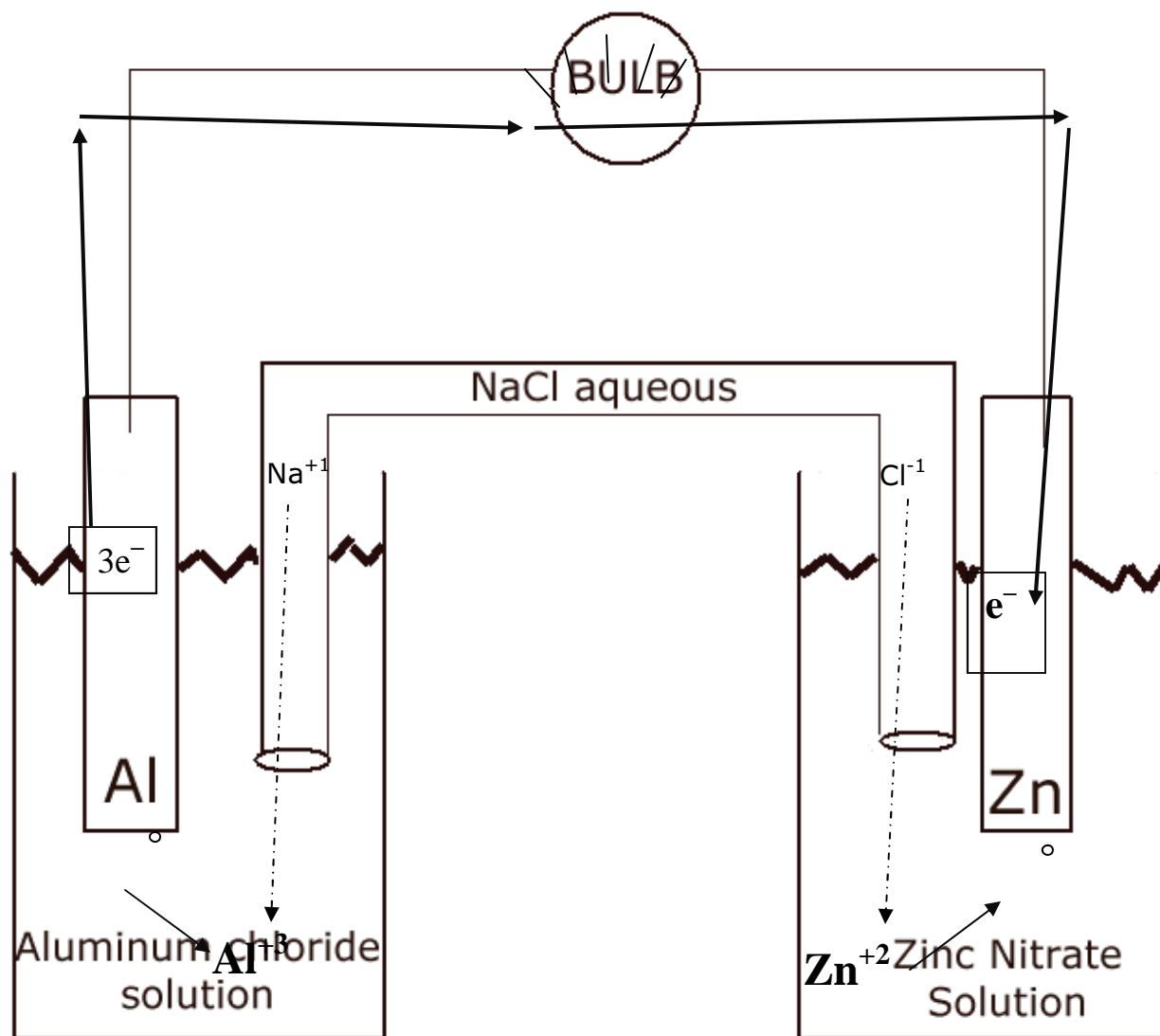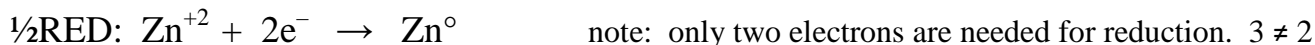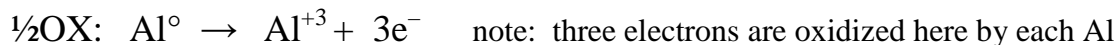
Each metal bar is an electrode. The names are cathode and anode. The way to keep them straight is to re-member Leo the Lion. He's a RED CAT. That should remind you that reduction happens on the cathode. Oxidation happens on the anode.

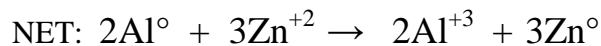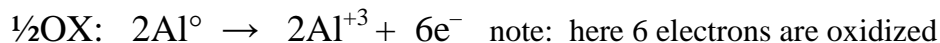So, aluminum is the anode, because the zinc is the cathode.

3 reasons that ALL batteries die…

1. Run out of anode. Sooner or later the aluminum atoms will all give up their electrons and jump into solu-tion, and there will be no atoms left, the bar will be dissolved.

2. Run out of salt bridge ions. Sooner or later all of the salt ions will all move to opposite beakers to keep the beakers neutral (offset the electrical potential). When that occurs, the solutions immediately become charged, stopping the flow of electricity.

3. Run out of cathode side cations. (NOT run out of cathode, that NEVER happens). If enough electrons arrive on the zinc bar, all the zinc cations in the solution will have become zinc atoms. At that point no more reduction can occur. The cathode gets bigger as the battery runs. Battery recyclers recycle the cathodes. Cathodes ARE NOT used up, they get bigger. Anodes get used up (see #1 above).

Let's quickly look at the REDOX of this battery. The two half reactions and the Net Ionic Equation.

$$\tfrac{1}{2}OX:\ Al° \ \rightarrow \ Al^{+3} + \ 3e^- \quad \text{note: three electrons are oxidized here by each Al}$$

$$\tfrac{1}{2}RED:\ Zn^{+2} + \ 2e^- \ \rightarrow \ Zn° \quad \text{note: only two electrons are needed for reduction. } 3 \neq 2$$

Let's balance using the LOWEST COMMON FACTOR (which is 6)

$$\tfrac{1}{2}OX:\ 2Al° \ \rightarrow \ 2Al^{+3} + \ 6e^- \quad \text{note: here 6 electrons are oxidized}$$

$$\tfrac{1}{2}RED:\ 3Zn^{+2} + \ 6e^- \ \rightarrow \ 3Zn° \quad \text{note: here, 6 electrons are reduced}$$

$$NET:\ 2Al° \ + \ 3Zn^{+2} \rightarrow \ 2Al^{+3} \ + \ 3Zn°$$

LOOK HARD: see that the charge is balanced in this net ionic equation, and that the six electrons oxidized are cancelled out by the six electrons that are being reduced. The electrons are omitted in the net ionic equation.

Charge must be conserved (the number of electrons are are oxidized must be perfectly balanced by the number of electrons being reduced. If it ever doesn't work, you are mixing up the charges of the ions being used. Check your details.

Electrolytic cells

In these cells, we use electricity to force redox, and we can use this to electoplate valuable metals onto less valuable metals. This makes our cheap filler be covered with precious metals, rings look great even though they might be inexpensive copper just coated with silver (for example).

These are NON-SPONTANEOUS reactions, but they are still redox. Below is an example diagram where we will "plate" silver onto a copper ring. Copper is cheap and strong but turns your finger green. Coating it in silver makes it look silver, makes you look richer, and keeps the green in the garden! Plus, copper insides are much stronger than a pure silver ring.

The electrons provided by the battery are pushed onto the copper ring. The silver ions in solution are attracted and get reduced to silver atoms, which plate on the copper. The longer this "runs" the thicker the silver coating gets.

To complete the circuit (and oxidize electrons to balance the reduction) the silver atoms in the silver bar oxidize, the silver cations replace the cations in solution, the electrons they oxidize run up the wire to the other side of the battery (replacing the electrons oxidized to provide the electricity in the first place that gets pushed onto the ring.
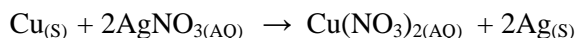
The ½OX: $Ag° \rightarrow Ag^{+1} + 1e^-$
this happens ON the silver bar

The ½RED: $Ag^{+1} + 1e^- \rightarrow Ag°$
this happens ON the ring

Electroplating is a billion dollar a year business. It is possible due to chemistry.

This is not a spontaneous redox, it requires energy, and that energy makes chemistry happen.

This is what happens when Copper is put into a silver nitrate solution:

$Cu_{(S)} + 2AgNO_{3(AQ)} \rightarrow Cu(NO_3)_{2(AQ)} + 2Ag_{(S)}$
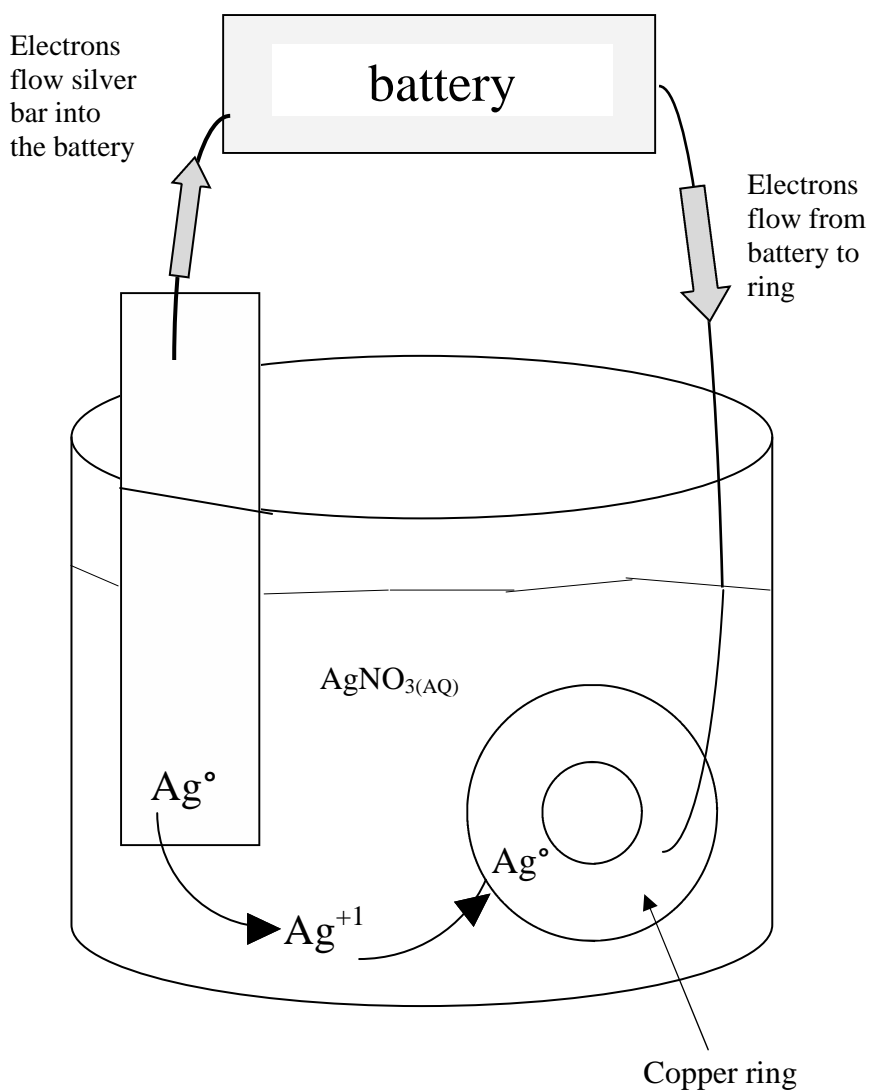
This is single replacement, and if you LET THIS HAPPEN, the silver precipitates spontaneously.

An electrolytic cell IS NOT SPONTANEOUS. You have to be smart, and you have to USE electricity to force a chemical reaction that you want to happen.

In a voltaic cell (a battery) the redox *is* spontaneous.

Redox is the name of a thrash metal band from Barcelona, Spain. There are apparently no fun photos or jokes to tell about redox.

Thanks for reading.

Electrons flow silver bar into the battery

battery

Electrons flow from battery to ring

$AgNO_{3(AQ)}$

$Ag°$

$Ag°$

Ag^{+1}

Copper ring

REDOX class notes

Redox is the part of chemistry that explains the chemistry behind
1. _____ 2. _____

3. _____

4. The reactions are always _____ (one is RED and one is OX)

5. In the old days, Oxidation meant _____

   and Reduction meant: _____

6. Now we understand that Oxidation means _____

   and Reduction means: _____

7. To remember this we will say OUT LOUD: _____ goes _____

8. Magnesium and sulfur make magnesium sulfide. Write the balanced chemical equation below

   _____ → _____

9. What do those little circles mean that you drew next the reactants above? _____

10. The product, magnesium sulfide is neutral, but what are the ion charges that SUM to zero? _____

11. In this reaction 2 different things happened:

    A. _____

    B. _____

12. Redox is
    _____

    _____

    _____

14. _____

15. Neutral atoms form a neutral ionic compounds, but the _____ numbers change.

16. Word equation: Silver nitrate and copper metal yields _____

17. Balanced chemical equation silver nitrate solution and copper yields copper (I) nitrate and silver metal.

_____

18. Add in the oxidation numbers (ionic charges) for the reaction above.

19. The Silver Ions _____ electrons, the Ag^{+1} are _____

20. The Copper atoms _____ electrons, the $Cu°$ are _____

21. The nitrate ions are still swimming around in the beaker. They are the _____ ions.

22. $Li_{(S)} + NaCl_{(AQ)} \rightarrow LiCl_{(AQ)} + Na_{(S)}$ is a single replacement/redox reaction. Write the 2 half reactions

½OX: _____

½OX: _____

NET: _____

23. _____

_____. The spectator ion is _____

24. $Mg_{(S)} + 2HCl_{(AQ)} \rightarrow MgCl_{2(AQ)} + H_{2(G)}$ is a SR/ redox reaction. Write it below as 2 half reactions

½OX: _____

½OX: _____

NET: _____

25. Here, the Mg atoms are oxidized into _____. The H^{+1} cations are reduced to _____

26. In this case, ____ electrons are oxidized, so ____ electrons must be reduced.

   The electron transfer MUST be in balance!

27. Write the balanced chemical reaction for this word equation:
   Sodium atoms + chlorine molecules synthesize into table salt

_____

   is a synthesis/redox reaction. Write it below as 2 half reactions

   ½OX: _____

   ½OX: _____

   NET: _____

_____

28. Oxidation numbers for ions are easy; they are just the _____

29. The oxidation number for the sodium cation is _____, For the chloride anion it is _____,

     for the sulfate anion (table E), it's _____, for magnesium cation it is _____.

30. For all atoms (+ the HONClBrIF twins) the oxidation number is _____, because they have no charge.

31. Inside molecules, like carbon dioxide (no ions) there are still oxidation numbers.
     What are the individual oxidation numbers for all of these species? (they better sum to zero!)

CO_2 _____   CO _____

$CaCl_2$ _____ NO_2 _____
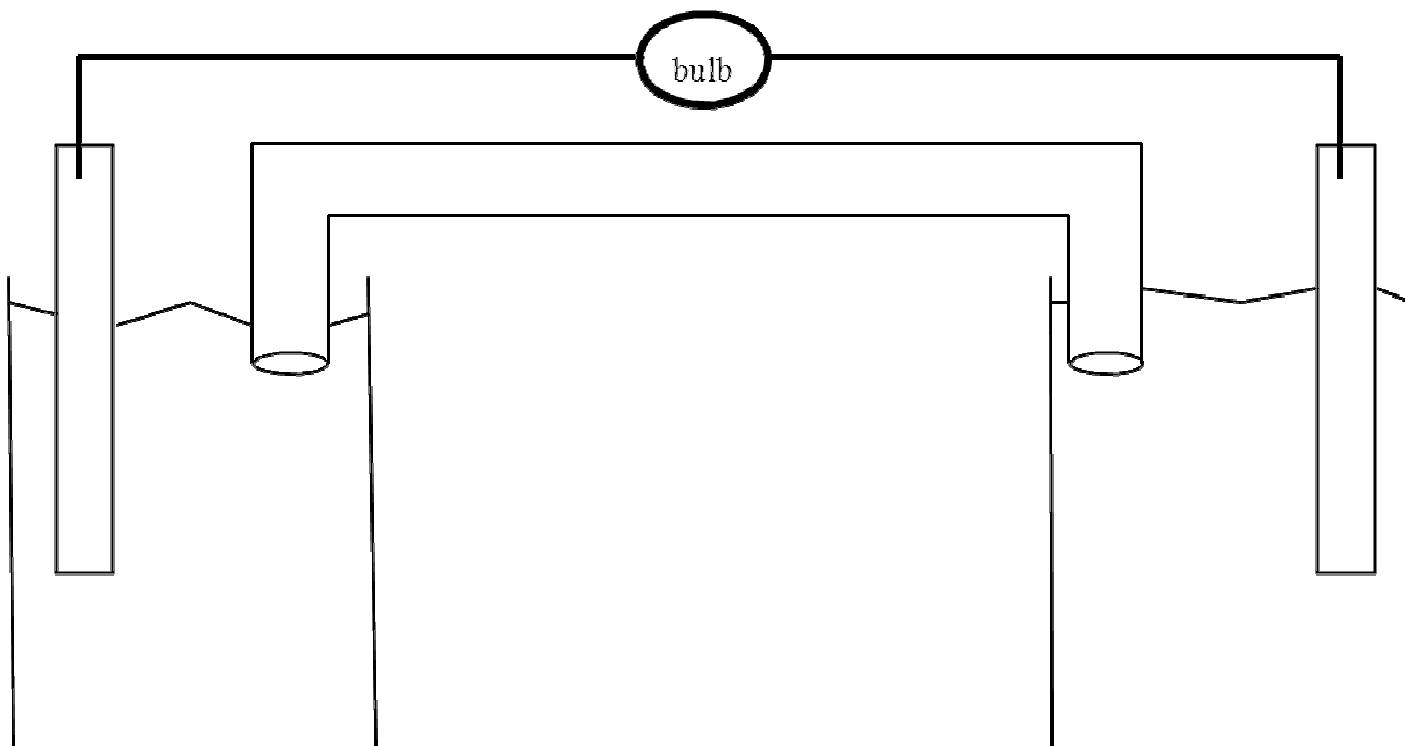
PCl_3 _____ PCl_5 _____

H_2SO_4 _____ $Cr_2O_7^{-2}$ _____ (special)

$NbBr_5$ _____

½OX: _____

½RED: _____

NET: _____

40. How do we decide which of the metals here will oxidize using table J?

41. Oxidation releases electrons, creating cations that jump into solution. Cations are positively charged.
These cations jump into a neutral solution, making the solution become _____ charged.

42. This solution, now positively charged, will attract ions from the salt bridge, which will neutralize the solu-
tion. A positive solution attracts _____ ions.

43. In the opposite beaker, where cations are forming into atoms on the cathode, the neutral solution is losing
the positive cations, making that solution become negatively charged. This negative solution will attract
the _____ salt ions to neutralize the solution.

46

bulb

NaCl$_{(AQ)}$

Salt bridge

Pb

Mg

PbCl$_{2(AQ)}$

Mg(OH)$_{2(AQ)}$

½OX:
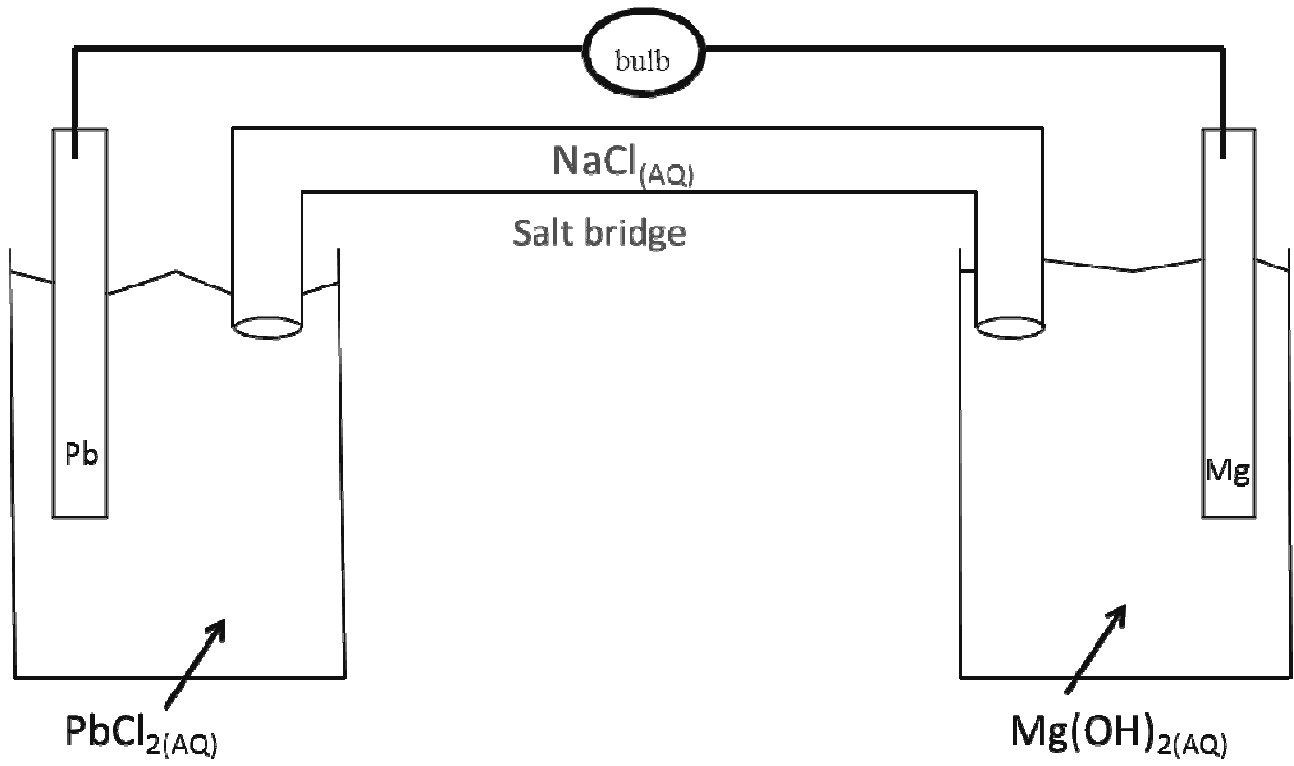
½RED:
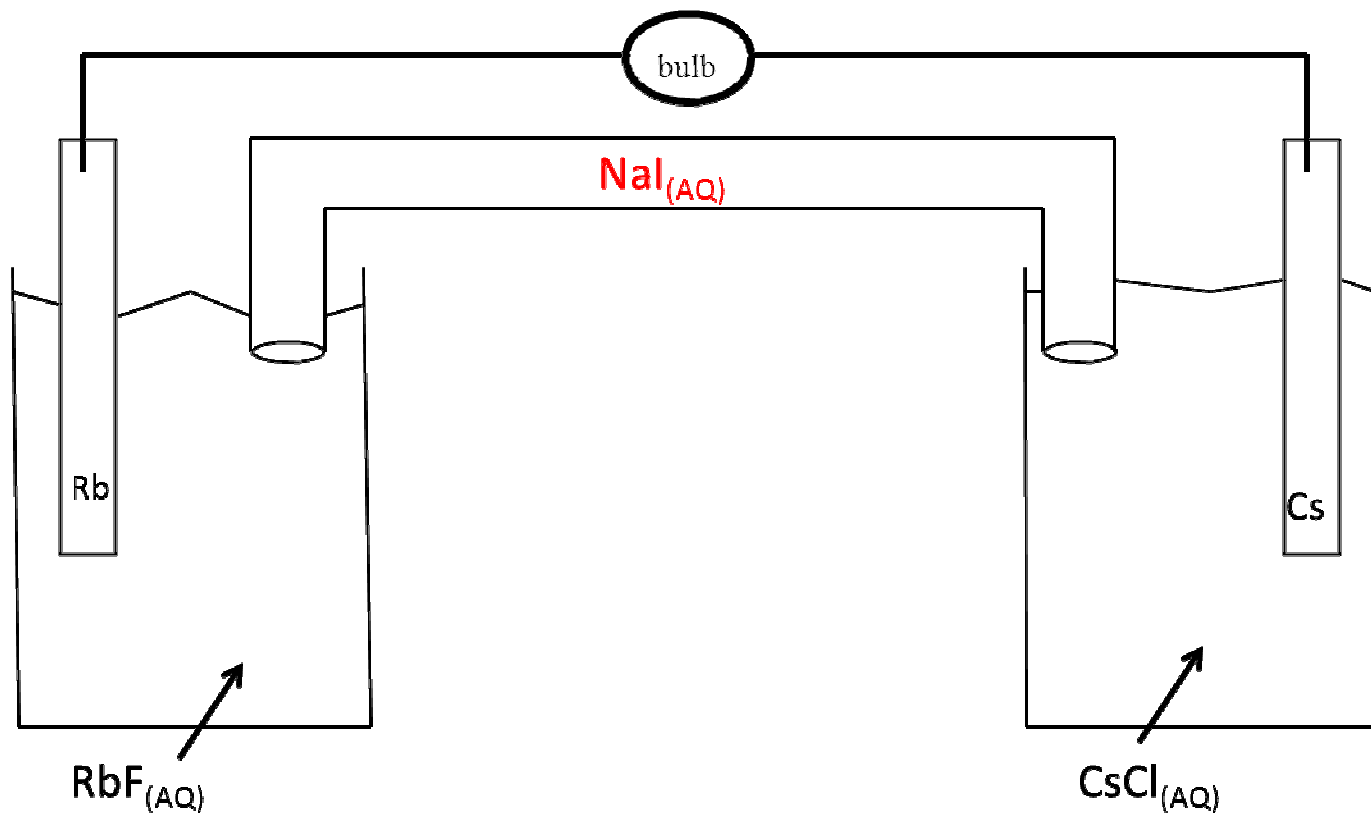
NET:

47. What are the 3 reasons that EVERY SINGLE battery dies, including this one?

A.

B.

C.

48. Completely Label this voltaic cell

bulb

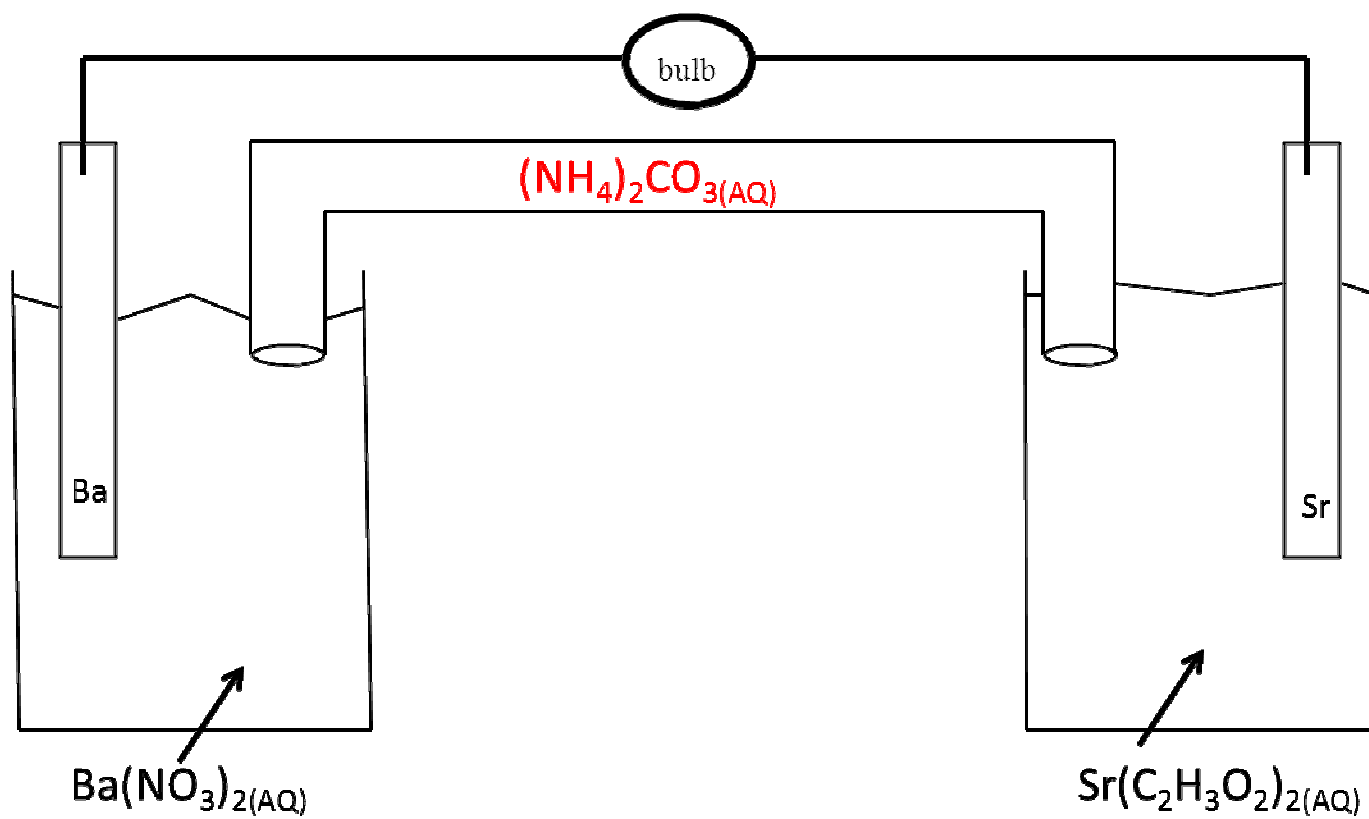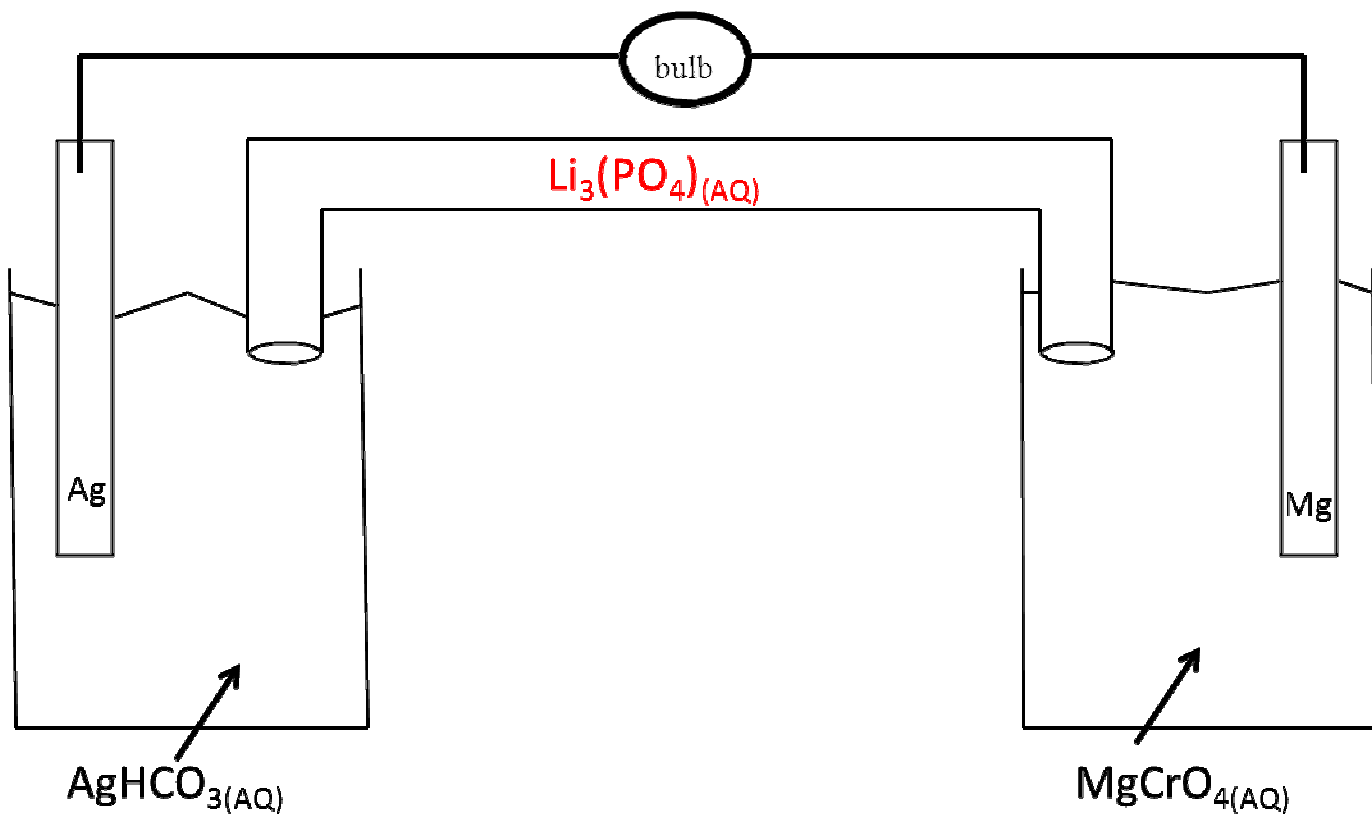$NaI_{(AQ)}$

Rb

Cs

$RbF_{(AQ)}$

$CsCl_{(AQ)}$

½ Oxidation: _____

½ Reduction: _____

Net Ionic Equation: _____

50. State the 3 specific reasons that **THIS voltaic cell** will die.

bulb

$(NH_4)_2CO_{3(AQ)}$

Ba

Sr

$Ba(NO_3)_{2(AQ)}$

$Sr(C_2H_3O_2)_{2(AQ)}$

½ Oxidation: _____

½ Reduction: _____

Net Ionic Equation: _____

53. State the 3 specific reasons that <u>THIS voltaic cell</u> will die.

bulb

$Li_3(PO_4)_{(AQ)}$

Ag

Mg

$AgHCO_{3(AQ)}$

$MgCrO_{4(AQ)}$

54. Write both half reactions, then the net ionic equation.

½ Oxidation: _____

½ Reduction: _____

Net Ionic Equation: _____

55. State the 3 specific reasons that <u>THIS voltaic cell</u> will die. Run out of

56. There are two kinds of ELECTROCHEMICAL CELLS, the one we know, Voltaic cells, and a new one

called the _____ Cell.

57. Voltaic cells have chemistry spontaneously creating electricity. An electrolytic cell REQUIRES

_____ to force a chemical reaction.

58. There are 2 kinds of electrochemical cells, the _____ + the

_____ cell.

59. When copper goes into silver nitrate solution, write out the spontaneous, balanced chemical equation.

_____
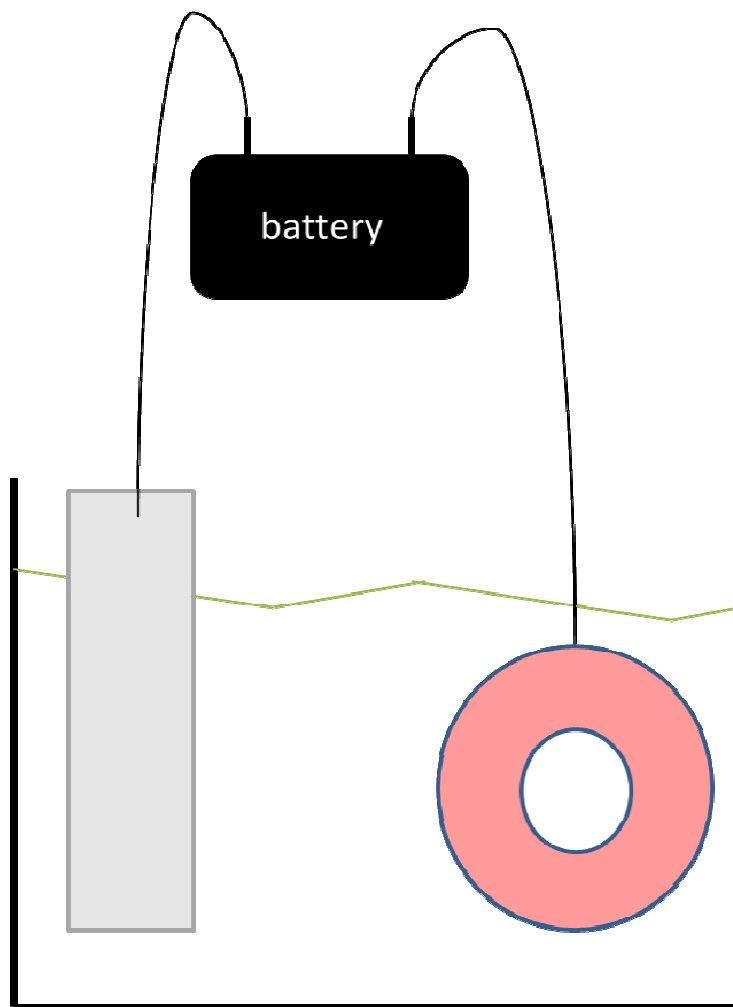
60. We can STOP THAT from happening, if we use an outside energy source, like a battery.

61. By using electricity to PUSH a nonspontaneous redox
   reaction, we've created an

   _____ cell

62. Label this diagram completely:
   flow of electricity, anode, cathode,
   state clearly what metal is going to
   electroplate onto the other.

battery

66. Write the half reactions now,
   try the net ionic equation!

   ½OX: _____

   ½OX: _____

   NET: _____

69. Draw the electrolytic cell that plates gold metal onto an aluminum spoon. Label it completely.

70. Write both half reactions, then the net ionic equation.

     ½ Oxidation: _____

     ½ Reduction: _____

     Net Ionic Equation: _____

73. Electrolysis:

   Hydrolysis:

74.

75. Write the electrical decomposition of water chemical equation, include the oxidation numbers

_____

76. Write out the half reactions for oxidation and reduction now.

   ½ Oxidation: _____

   ½ Reduction: _____

   NET IONIC EQ: _____

Quick Review…

77. Name the type of electrochemical cell that spontaneously produces electricity from a chemical reaction.

   _____cell

78. Name the type of cell where electricity forces a redox reaction that would not be spontaneous.

   _____cell

79. What always happens at the anode? _____

80. What always happens at the cathode? _____

81. Is Leo ALWAYS a RED-CAT?

Some of the 100 Redox Practice Problems, which are online waiting for you!

Redox/Electrochemistry Regents Unit Review.   In chemistry we will sometimes see ions written as Zn^{+2} but sometimes as Zn^{2+}. I mixed them up here so you don't get "surprised". Sometimes Na^{+1} will be written as Na^+, without the "one". Cl^{-1} might be "just" Cl^-. It's okay, relax, you're smart.

Base your answers on the following redox reaction, which occurs spontaneously in an electrochemical cell:

$$Zn° + Cr^{3+} \rightarrow Zn^{2+} + Cr°$$
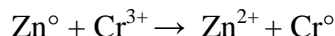
1. Write the half-reaction for the reduction that occurs.
2. Write the half-reaction for the oxidation that occurs.
3. Balance the equation using the smallest whole-number coefficients.
4. Which species loses electrons and which species gains electrons?
5. Which half-reaction occurs at the cathode?
6. What happens to the number of protons in a Zn atom when it changes to Zn^{2+} in the redox reaction.

Given the reaction: $Mg_{(S)} + 2H^+_{(AQ)} + 2Cl^-_{(AQ)} \rightarrow Mg^{2+}_{(AQ)} + 2Cl^-_{(AQ)} + H_{2(G)}$

7. Which species undergoes oxidation?   A. $Mg_{(S)}$ B. $H^+_{(AQ)}$ C. $Cl^-_{(AQ)}$ D. $H_{2(G)}$

8. Which particles are gained and lost during a redox reaction?
   A. electrons   B. neutrons   C. protons   D. positrons

9. As a Ca atom undergoes oxidation to Ca^{2+}, the number of neutrons in its nucleus
   A. increases   B. decreases   C. remains the same   D. no one knows this, me either
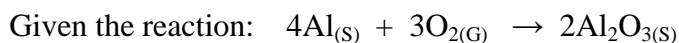
Given the reaction: $4Al_{(S)} + 3O_{2(G)} \rightarrow 2Al_2O_{3(S)}$
10. Write the balanced oxidation half-reaction for this oxidation-reduction reaction.
11. What is the oxidation number of oxygen in Al_2O_3?

12. In any redox reaction, the substance that undergoes reduction will
    A. lose electrons and have a decrease in oxidation number
    B. gain electrons and have a decrease in oxidation number
    C. lose electrons and have an increase in oxidation number
    D. gain electrons and have an increase in oxidation number

14. When a neutral atom undergoes oxidation, the atom's oxidation state
    A. decreases as it gains electrons   B. decreases as it loses electrons
    C. increases as it gains electrons   D. increases as it loses electrons

15. Given the equation: $C_{(S)} + H_2O_{(G)} \rightarrow CO_{(G)} + H_{2(G)}$ Which species undergoes reduction?
    A. $C_{(S)}$ B. H^+ C. C^{2+} D. $H_{2(G)}$

16. Given the equation: $2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$ The reduction half-reaction is...
    A. $Al \rightarrow Al^{3+} + 3e^-$ B. $Al + 3e^- \rightarrow Al^{3+}$ C. $Cu^{2+} + 2e^- \rightarrow Cu$ D. $Cu^{2+} \rightarrow Cu + 2e^-$

17. Which type of reaction occurs when nonmetal atoms become negative nonmetal ions?
    A. oxidation   B. substitution   C. reduction   D. condensation

18. Given the reaction: $Zn_{(S)} + 2HCl_{(AQ)} \rightarrow ZnCl_{2(AQ)} + H_{2(G)}$

   Which statement correctly describes what occurs when this reaction takes place in a closed system?
   A. Atoms of $Zn_{(S)}$ lose electrons and are oxidized.   B. Atoms of $Zn_{(S)}$ gain electrons and are reduced.
   C. There is a net loss of mass.   D. There is a net gain of mass.

19. Given the reaction: $Cl_2 + 2HBr \rightarrow Br_2 + 2HCl$
   Write a correctly balanced reduction half-reaction for this equation.

20. Given the reaction: $2 Al_{(S)} + Fe_2O_{3(S)} \rightarrow Al_2O_{3(S)} + 2Fe_{(S)}$
   Which species undergoes reduction?   A. Al   B. Al^{3+} C. Fe^{3+} D. Fe

21. Which equation shows conservation of both mass and charge?
   A. $Cl_2 + Br \rightarrow Cl^- + Br_2$ B. $Cu + 2Ag^+ \rightarrow Cu^{2+} + Ag$
   C. $Zn + Cr^{3+} \rightarrow Zn^{2+} + Cr$ D. $Ni + Pb^{2+} \rightarrow Ni^{2+} + Pb$

22. Given the reaction that occurs in an chemical cell: $Zn_{(S)} + CuSO_{4(AQ)} \rightarrow ZnSO_{4(AQ)} + Cu_{(S)}$
   During this reaction, the oxidation number of Zn changes from
   A. 0 to +2   B. 0 to -2   C. +2 to 0   D. -2 to 0

23. Given the reaction for the corrosion of aluminum: $4Al + 3O_2 \rightarrow 2Al_2O_{3(S)}$
   Which half-reaction correctly represents the oxidation that occurs?
   A. $Al + 3e^- \rightarrow Al^{3+}$ B. $Al \rightarrow Al^{3+} + 3e^-$ C. $O_2 + 4e^- \rightarrow 2O^{2-}$ D. $O_2 \rightarrow 2O^{2-} + 4e^-$

24. Balance this *unbalanced* redox reaction: $Cu_{(S)} + AgNO_{3(AQ)} \rightarrow Cu(NO_3)_{2(AQ)} + Ag_{(S)}$
25. Write the reduction half reaction from it now.
26. Besides redox, what kind of chemical reaction is the reaction in question 24?

27. Which half-reaction correctly represents reduction?
   A. $Ag \rightarrow Ag^+ + e^-$ B. $F_2 \rightarrow 2F^- + 2e^-$ C. $Au^{3+} + 3e^- \rightarrow Au$ D. $Fe^{2+} + e^- \rightarrow Fe^{+3}$

The rest of the questions (there are 100) and all of the answers are at
www.Arbuiso.com on the REDOX page.

This is called the 100 Redox Questions.