

CHM 101: GENERAL CHEMISTRY I

TOPIC: Redox Reaction and Introduction To Electrochemistry

LECTURER: PROF. OLOWU R.A

Learning Objectives:

At the end of this lesson, the students should be able to:

1. Define oxidation and reduction
2. Calculate the oxidation number of simple and complex compounds
3. Identify the oxidizing agent and reducing agent in a redox reaction
4. Construct an electrochemical cell and calculate the cell potential (emf)
5. Determine the standard reduction potential and use its value to predict the feasibility of a reaction
6. Determine Gibb's free energy and equilibrium constant from standard reduction potential
7. State the applications of cell potential
8. Construct an electrolytic cell and write a balanced net equation in the cell
9. Determine the mass/volume of a substance deposited at the electrode during electrolysis.

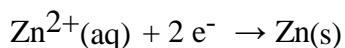
GENERAL REVIEW

Electrochemistry is the interchange of chemical and electrical energy. The key to electrochemistry is a redox reaction. This unique type of reaction produces a flow of electrons that can be used to do work like light a flashlight bulb or perhaps your favorite...provide energy for your cell phone to function. As you might notice the first type of electrochemical cell that we will discuss is a battery or in scientific terms, a galvanic cell. Before we can discuss the details of a battery, background information about redox reactions needs to be understood.

REDOX

The term “redox” is really a short way of referring to an oxidation-reduction reaction. **Oxidation** occurs when a species (an atom or ion) loses electrons in a reaction. **Reduction** occurs when a different atom or ion in the reaction gains the electrons lost. The acronym ‘**OIL RIG**’ (Oxidation Involves Loss of electrons; Reduction Involves Gain of electrons) *OR LEO the lion says GER* (Losing Electrons is Oxidation; Gaining Electrons is Reduction) is used in describing redox.

Note that since electrons are negative, if reduction occurs then the oxidation number of the species is reduced or will go down in value. For example,



Notice that the Zn^{2+} ion **gained** 2 electrons (electrons are on the reactant side) so it's charge or oxidation number went down to zero as solid Zn has no overall charge (oxidation number is

zero).

If oxidation is occurring, electrons are **lost** so the electrons are on the product side of the reaction, and the oxidation number of the product should increase. For example,



Solid Ni has an oxidation number of 0 and increases to +2 when it loses 2 electrons to become Ni^{2+} .

Notice that the law of conservation of mass is obeyed. Now we not only need to make sure that the numbers of atoms are balanced on both sides of the reaction, but also to confirm that both sides of the reaction have charges balanced. Notice in the zinc example that there is one Zn and on each side of the yield sign AND that the reactant side = zero since $+2 + -2 = 0$ just as solid zinc on the product side is zero or has no charge.

OXIDATION NUMBERS

Use the following **rules to determine oxidation numbers** so that we can distinguish between which atoms are being oxidized and reduced based on how these numbers are changing in a redox reaction.

1. Elements have an oxidation number = zero (Fe, Cu, Na, O₂, H₂, P₄, etc)
2. Monoatomic ions equals the ion's charge (Cu^{2+} , Cu^+ , Ag^+ , Al^{3+} , Cl^- , S^{2-} , etc)
3. Fluoride is -1
4. Oxygen in a compound or ion is -2 except in peroxide (O_2^{2-}) where each O is -1 and if O is bonded to F (OF₂) then O is +2
5. Hydrogen is +1 or if with a metal is -1 (H^{-1} is called hydride)
6. Oxidation numbers in compounds must equal zero.
7. Oxidation numbers in a polyatomic ion must equal the charge of the ion.
8. Assign the most electronegative element a negative oxidation number

Examples-

1. **H₂O** is a compound (rule #6) so the total of oxidation numbers = 0. Since O is usually -2, then each H must be +1 for the total to equal zero.
2. **MnO₄⁻** is a polyatomic ion (rule #7) so the total of all oxidation numbers must equal the overall charge of -1. (Always start with oxygen if it's present.) If each O is -2 (that's 4 times -2 = -8) than Mn must be +7 so that $+7 + -8 = -1$
3. What about **H₂O₂**? It looks similar to H₂O but is hydrogen peroxide so be

careful. Each O would only have an oxidation number of -1 and each H is +1 so that $(+1 \times 2) + (-1 \times 2) = 0$

4. **KClO₃** is a compound so the total of oxidation numbers is 0. When a compound is given that contains a polyatomic ion, consider splitting it into ions so that assigning oxidation numbers is easier.

- a. KClO₃ is really a combination of K⁺¹ and ClO₃⁻¹ to equal zero. So the oxidation number of K is +1.
- b. Now Chlorate ion can be considered. O is -2 and since there are a total of 3 (3 \times -2), the overall charge of all O is -6. Cl would have to be +5 so that $+5 + -6 = -1$.

Check to make sure all oxidation numbers equal zero for the whole compound. K is +1 added to +5 for Cl added to 3 O at -2 each (-6 total) does equal zero

Balancing Redox Reactions

Since redox reactions are unique in that electrons are transferred between oxidized and reduced substances, the balancing is unique as well. This type of balancing is a process. Remember these 4 items to balance in this order:

1. Balance any atom other than O or H.
2. Balance O by adding water.
3. Balance H by adding H⁺.
4. Balance charge by adding electrons.

Sounds weird, right? Many redox reactions are very simple and not complicated so don't freak out.

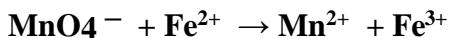
Steps in balancing redox reactions

1. Split the reaction into 2 half reactions. One half will show oxidation and one half will show reduction. Each half will be balanced separately and then at the end the halves will be added back together to get an overall total reaction.
2. Starting with one of the half reactions, balance all atoms other than O and H using coefficients.
3. Balance O by adding water to either side of equation as needed.
4. Balance H by adding H⁺ ion to either side of equation as needed.
5. Add up all charges on each side of the half reaction. Add as many electrons as needed to one side so that charges of reactants now equal charges on products side.
6. Go to the other half reaction and repeat steps #2-5.
7. Look at the electrons in both half reactions. The number of electrons lost must equal the number of electrons gained. If the number of electrons in each half reaction is

not equal then multiply the entire half reaction by an integer that will make the electrons balanced. (You might need to multiply both half reactions by a whole number so that electrons are equal.)

8. Add the half reactions back together and the electrons should cancel since you should have the same coefficient in front of the electrons lost (on product side) and the electrons gained (on reactant side).
9. Check to see that atoms and charges balance.

Balancing Redox Example #1 (longer balancing using all steps) Balance the following reaction.



You can tell this is unusual since O is missing from the product side. The redox balancing steps will remedy this issue.)

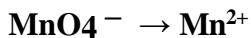
Fe^{2+} is oxidized since it's oxidation number goes from +2 to +3. This can be the first half reaction to balance.

$\text{Fe}^{2+} \rightarrow \text{Fe}^{3+}$ Step 2- Fe is balanced already; Step 3- no O to balance; step 4- no H to balance

Go to step 5 and balance charges by adding electrons. If 1 e^- is added to the product side then both sides have the same charge of +2.

** $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + 1 e^-$ (this is balanced- notice oxidation MUST have the electron as a product since it LOSES electrons)

Other half reaction:

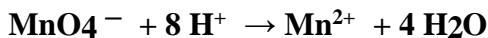


step 2- Mn is balanced

step 3- need to balance O. 4 on the reactant side so add 4 water molecules to the product side. Now there are 4 O on both sides of the equation.

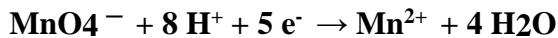


step 4- need to balance the 8 H (from water) on the product side by adding 8 H^+ to the reactants.



step 5- add up charges on each side of the equation.

-1 + **+8** **+2** + **0** +7 is not equal to +2 so 5 e⁻ are needed on the reactant side for both reactants and products to have an overall charge of +2.

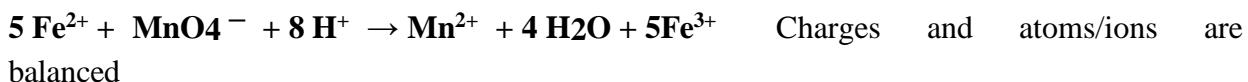


This is now balanced and shows a reduction since electrons are gained.

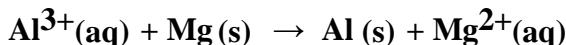
Look at the 2 starred and balanced half-reactions. Are there equal numbers of electrons that are lost and gained in each? No, so lastly electrons must be balanced. The Fe half-reaction will be multiplied by 5 to balance the electrons.



The 2 half reactions are now balanced, electrons are equal, and cancel, add the reactions.



Balancing Redox Example #2



Go back to the balancing steps. Step 1-split the reaction into halves.

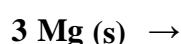
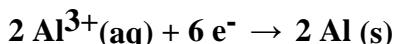
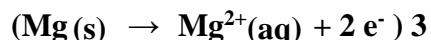


Notice that there is no oxygen present in either half reaction so the balancing will be much simpler. Step 2 Both Al and Mg are balanced; Step 3- No O to balance; Step 4- No H to balance;

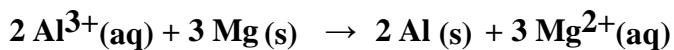
Step 5- will always be needed since electrons are being transferred. Look at charges on the reactant and product sides of each half-reaction. Determine how many electrons are needed in balancing.



Step 7- the electrons in each half-reaction need to be equal. To balance 2 and 3 electrons, multiply each half-reaction by the correct integer to get 6 electrons.



3 Mg²⁺(aq) + 6 e⁻ Step 8- Add the 2 half reactions together and the electrons cancel.



Charges and atoms/ions are balanced.

Galvanic Cells

Now that you understand more about what redox reactions are and how to balance them, we can discuss their application as electrochemical cells. As mentioned earlier, a battery is one type of an electrochemical cell called a **galvanic cell** (or voltaic cell). Think of batteries that are important in everyday life. The first to come to mind might be your cell phone battery or the batteries in your calculator. How many batteries can you come up with? (car batteries, computer, AA, 9-volt, etc.) Every battery can have different components, but the commonality is that all are chemical reactions that produce electrical energy we can use.

2 Types of Electrochemical Cells:

1. **Galvanic Cells** (a battery)
2. **Electrolytic Cells** (which require a power source) will be discussed later

Galvanic Cell

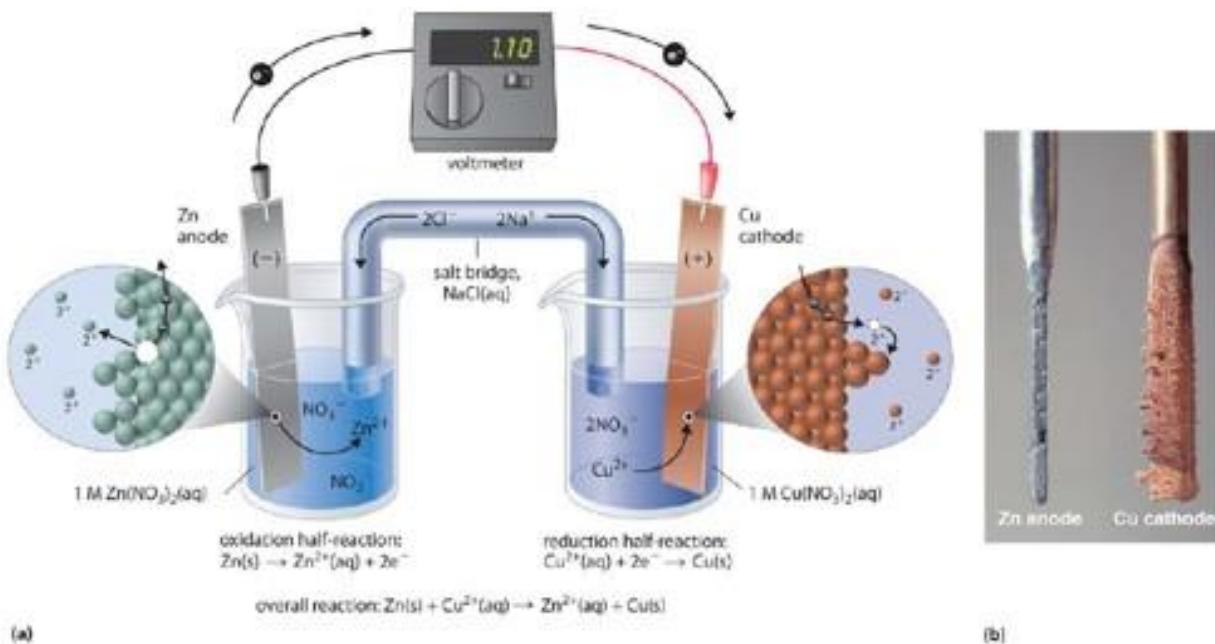


Figure 1: Galvanic Cell

Components of Galvanic cell

1. **Anode**- the electrode where oxidation occurs. (**An Ox anode** is where oxidation occurs). Mass can decrease
2. **Cathode**- the electrode where reduction occurs. (**Red Cat reduction** occurs at the cathode) Mass can increase.
3. **Voltmeter**- measures the cell potential of the battery in Volts (V). A wire connects the anode and cathode to the voltmeter. The electrons leave the anode and travel through the wire to the cathode. (**FAT CAT – Flow of electrons is from Anode To Cathode**)
4. **Salt bridge**- connection between the anode and cathode that maintains electrical neutrality by providing cations and anions to keep charges balanced. Cations (+ ions) from the salt bridge flow to the cathode and anions (- ions) from the salt bridge flow to the anode.

Examples: piece of filter paper soaked in a neutral salt, electrolyte paste, or a porous cup. The salt bridge is generally composed of a dissolved salt containing ions that will not react with other components in the galvanic cell (like KNO₃). Think spectator ions.

How a Galvanic Cell operates (refer to figure 1a)

- Notice that all components (reactants and products) of the reaction are present before the reaction begins.
- If solid metals are part of the reaction, then the electrodes are made of those metals. If only ions are present in the reaction (no solid metals) then the electrodes are made of inert (or an unreactive) substance like graphite (a form of carbon i.e. pencil “lead”) or Platinum (Pt).
- In Figure 1, Zinc metal (the electrode) is dipping into a solution containing dissolved Zn(NO₃)₂ solution as determined by the presence of separate Zn²⁺ and NO₃⁻ ions, and Cu metal (the electrode) is dipping into a solution of dissolved Cu(NO₃)₂ as determined by the presence of separate Cu²⁺ and NO₃⁻ ions. The NO₃⁻ ions will not be part of the reaction since they are spectator ions and stay dissolved throughout the reaction.
 - **A metal dipping into a solution that contains its own ion will indicate the half reaction.**
- The electrodes have previously been labeled anode (Zn) and cathode (Cu) along with the electrons shown are traveling through the wire from the Zinc cell to the Copper cell (anode to cathode).
- How can we determine and label in future galvanic cells which half reactions should be oxidation or reduction? With each electrochemistry question a table of reduction potentials will be provided. These values (voltages) indicate how likely a substance is to be reduced. Compare the values for the 2 half reactions in a galvanic cell. The more positive potential indicates the substance that will be reduced so we can deduce that the other substance with the lower reduction potential must be oxidized since reduction

cannot occur without oxidation. See the table below:

Standard Reduction Potentials in Aqueous Solutions at 25°C

Half-reaction	E° (V)
$\text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu}$	0.34
$\text{Zn}^{2+} + 2 \text{e}^- \rightarrow \text{Zn}$	-0.76

- In Figure 1, Cu is reduced since its reduction potential above is more positive. Zn must be oxidized since it had the lower reduction potential. Notice that if a reduction half-reaction is reversed, it is oxidation.
 - Zn is not reduced so reverse its half-reaction. $\text{Zn} \rightarrow \text{Zn}^{2+} + 2 \text{e}^-$
 - This is the half-reaction indicated in figure 1.
 - What happens to a reduction potential if the reaction is reversed to show oxidation?
 - The sign is reversed as well.
 - The oxidation potential for Zn is +0.76V
- Also, in figure 1, the voltmeter has a reading of 1.10 (V). This shows us the maximum cell potential we could expect to attain from this reaction. Its calculation is simple.

$$E^\circ \text{ (standard cell potential)} = E^\circ \text{reduction} + E^\circ \text{oxidation}$$

For figure 1, $E^\circ = 0.34 + 0.76 = 1.10\text{V}$ as shown on the voltmeter.

- **Why is the Salt Bridge needed?**

- As Zn atoms lose electrons and Zn^{2+} goes into solution, there is a build-up of positive charges in the anode. Anions from the salt bridge (Cl^- in figure 1) are attracted to these extra positive charges which then neutralizes the positive charge build up allowing electrons to continue to flow.
- As Cu^{2+} ions gain electrons to form solid Cu (and plate on top of the existing Cu electrode), positive charges are missing from the solution which was previously balanced with the nitrate spectator ions. With positive charges leaving, the negative nitrate ions are in a greater amount so cations from the salt bridge (Na^+ in figure 1) are needed to flow into the cathode and neutralize the negative charges so that electrons can continue to flow.

Bottom line: No salt bridge, no electron flow, voltmeter would read ZERO

Questions you might have:

- What is happening In Figure 1b where the zinc anode appears to have been “eaten away” and the copper electrode looks larger?
 - This is what was meant in the earlier notes when the anode can have its mass decrease while the mass of the cathode can increase. When the electrodes are metals in the reaction, the electrode masses will change depending on which process occurs at its surface.
 - Where does the mass go for the zinc or where does the new mass plating on the copper electrode come from? Ions are the key. Solid zinc loses 2 electrons (electrons’ mass are so small it’s negligible) so the mass of the solid zinc electrode goes into the solution as Zn^{2+} . Conversely the Cu^{2+} gains 2 electrons and plates onto the existing copper electrode- zinc’s mass decreases and copper’s mass increases.
- Where do the reduction potentials values come from?
 - Each potential is measured against a standard. For the potentials we will use, the standard is hydrogen gas which is assigned a potential of zero volts. Half reactions with positive potentials are substances that will be reduced and hydrogen would be oxidized, and negative potentials are substances that will be oxidized when hydrogen is reduced.
- What is the “degree” symbol I see with the cell potential abbreviation, E° ?
 - This symbol is a “naught” sign, and denotes standard conditions being present.
 - Standard conditions are: $25^\circ C$ ($298K$), 1.0 M for solutions, and 1 atm for gases
- What is cell potential or voltage? **the measure of the potential difference between two half cells in an electrochemical cell.** The potential difference is caused by the ability of electrons to flow from one half cell to the other
 - Common symbols are E_{cell} or Emf o r E°_{cell}
 - Unit is the Volt (V)

The standard cell potential (E°_{cell}) is therefore the difference between the tabulated reduction potentials of the two half-reactions, not their sum:

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$

NOTE: Cell with higher potential represent the cathode and smaller one is the anode.

APPLICATION OF CELL POTENTIAL

1. To determine the direction of electron flow

2. To predict the feasibility of reactions
3. To determine the strength of reducing agent and oxidizing agents
4. To predict the standard cell potential, E°_{cell}

Gibb's Free Energy and Electrochemistry

Now that you can determine the anode, and cathode, and understand the purpose of components, let's expand into a little more detailed energy discussion.

Emf (V) = or rearrange the units: Joule = Volt x Coulomb or J = V x C

- The work that can be accomplished when electrons are transferred through a wire depends on the potential difference (in volts) between 2 points in the circuit.
- One Joule (J) of work is produced when one Coulomb (C) of charge is transferred between two points in the circuit that differ by a potential of one volt. [A Coulomb is a quantity of electrons or charge passing through the circuit]
- If work flows OUT it is assigned a minus sign so that when a galvanic cell produces electricity the cell potential is positive and can be used to do work so work has a negative sign.
- All of this background information is to help us understand:
 - $\Delta G^\circ = -nFE^\circ$
 - **G** is Gibb's Free Energy (discussed in great detail in Thermodynamics)
 - It's sign is very important to us. For a reaction to occur on its own or be **thermodynamically favorable** the sign of ΔG° must be **negative**.
 - Batteries are chemical reactions that occur when components are in contact with one other so are thermodynamically favorable ($\Delta G^\circ < 0$)
 - **n** is the number of moles of electrons transferred (balanced number from the half-reactions that cancels out)
 - **F** is Faraday's constant (the charge of one mole of electrons) = 96500 C/mole of electrons
 - **E°** is the standard cell potential from $E^\circ = E^\circ_{reduction} - E^\circ_{oxidation}$
- Going back to Figure 1a, the $E^\circ = 1.10$ V.
 - **What is the value of ΔG° ?**
 - $\Delta G^\circ = -nFE^\circ$ ($n = 2$ since both half-reactions each lost and gained 2 electrons)
 - $\Delta G^\circ = -(2 \text{ mol e}^-)(96500 \text{ C/mol e}^-)(1.10\text{V})$
 - $\Delta G^\circ = -212000 \text{ V}\cdot\text{C} = -212000 \text{ J or } -212 \text{ KJ}$.

Is this reaction thermodynamically favorable? Justify your answer.

Yes, ΔG° is negative since the cell potential is positive.

- A galvanic cell or battery is a chemical reaction that occurs on its own. We know that the reaction is going forward.
 - Another important variable is **K, the equilibrium constant**
 - K has no units (WOW, right?)
 - Describes how the reaction proceeds and is simply a ratio of .
 - If a reaction like a galvanic cell is going forward, it produces a lot of products and uses the reactants so that the ratio of $K > 1$.
 - The value of K can predict the sign of ΔG° as well.
 - $\Delta G^\circ = -RT \ln K$
 - **G** is Gibb's free energy (remember a negative ΔG° tells us the reaction is thermodynamically favorable)
 - **R** is the energy ideal gas constant (watch units; pick the one on the formula sheet with Joules) 8.314 J/K•mol
 - **T** is temperature in Kelvin
 - **ln** is natural logarithm (find this button on your calculator)
 - **K** is the equilibrium constant (no units)
- Going back to figure 1a and assume temperature is 25°C, **calculate the value of the equilibrium constant.**
 - **Above $\Delta G^\circ = -212267 \text{ J}$** (it's better to not round until the end of a problem so I'm using the unrounded value)

$$\Delta G^\circ = -RT \ln K$$

$$\ln K = \frac{-\Delta G^\circ}{RT} = \frac{-(-212267)}{8.314 \times 298} = 85.68$$

$$K = 1.62 \times 10^{85.68}$$

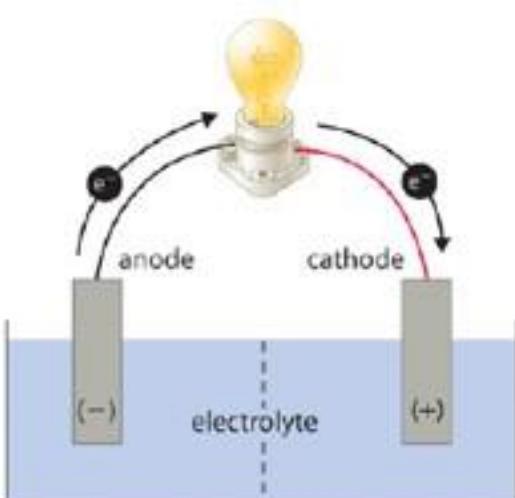
since $K \gg 1$, the reaction is definitely going forward.

Galvanic cells are batteries so
E° is positive
ΔG° is negative
$K > 1$

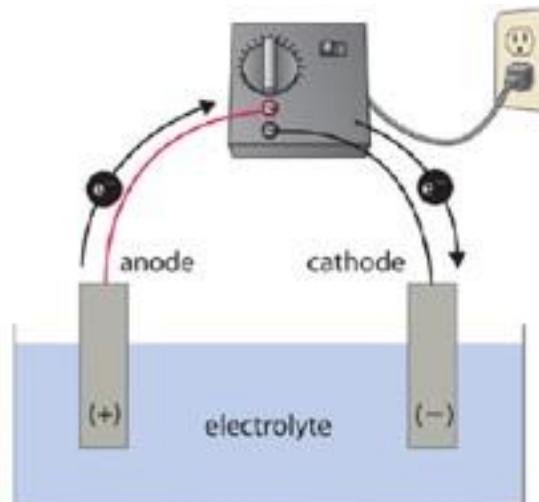
Galvanic Cells vs Electrolytic Cells

Carefully review how electrolytic cells are different from galvanic cells. Many students

consider them to be “opposite of each other



GALVANIC CELL



ELECTROLYTIC CELL

Similarities

Oxidation occurs at the anode and reduction occurs at the cathode	Oxidation occurs at the anode and reduction occurs at the cathode
Electrons flow from anode to cathode	Electrons flow from anode to cathode

Differences

Produces electricity (to light the bulb above)	Requires a power source
2 containers or 1 with an electrolyte barrier between the cells	All components are in one container
Has an electrolyte barrier (salt bridge)	Does not need a salt bridge
Cathode is positive	Cathode is negative
Most positive reduction potential stays reduction	Most positive reduction potential and half reaction are reversed to be oxidation
Lower reduction potential and half reaction are reversed to be oxidation	Lower reduction potential stays reduction
E° is positive	E° is negative
ΔG° is negative, thermodynamically favorable	ΔG° is positive, thermodynamically unfavorable
$K > 1$, products are favored	$K < 1$, reactants are favored

ELECTROLYTIC CELL

This is a cell where electrolysis takes place. The cell convert an electrical energy to non spontaneous reaction.

In bold, is a significant difference that subsequently causes the opposite signs for E° , ΔG° , and K. Notice that for electrolytic cells the higher half reaction potential is the one that will be reversed as oxidation and the lower potential will stay reduction. For example,

Half-reaction	E° (V)
$O_2(g) + 4 H^+(aq) + 4 e^- \rightarrow 2 H_2O(l)$	+1.23
$Cu^{2+}(aq) + 2 e^- \rightarrow Cu(s)$	+0.34

- Write a balanced net ionic equation for the electrolysis reactions that would occur in the cell

- Since this is an electrolytic cell, the Cu half-reaction with the lower potential will remain reduced, and the higher potential of the oxygen gas half-reaction will be reversed and its sign changed to negative to signify oxidation.
- Electrons need to be balanced.
- $(Cu^{2+}(aq) + 2 e^- \rightarrow Cu(s)) 2$
 - $2 H_2O(l) \rightarrow O_2(g) + 4 H^+(aq) + 4 e^-$
 - $2 H_2O(l) + 2 Cu^{2+}(aq) \rightarrow O_2(g) + 4 H^+(aq) + 2 Cu(s)$
(Check to verify that atoms and charges are balanced)

- Calculate the standard cell potential.

- Cu half reaction with the lower potential will remain reduction, **+0.34 V**
- Higher potential of the oxygen gas half reaction will be reversed and its sign changed to negative to signify oxidation, **-1.23 V**
- $E^\circ = +0.34 + -1.23 V = -0.89 V$

- Is the reaction thermodynamically favorable? Justify your answer.

- No, the reaction is not thermodynamically favorable since E° is negative causing G° to be negative.

- Is how many coulombs of charge (electrons) pass a given point per second
- Symbol is I
- Formula is $I = Q/t$
 - Current = quantity of charge/time
- Measured in Amperes (A)
or “amps”
- Amperes = Coulombs/seconds
 - $1\text{ A} = 1\text{ C/s}$
 - $Q = It$

USE $I = Q/t$ to answer these questions:

1. How long must a current be applied to a solution to plate out a metal?
2. What mass of metal will plate out if a current is applied a certain amount of time? These are questions can be answered by doing dimensional analysis.

Faraday First law of Electrolysis: This law state that the mass of a substance liberated during electrolysis is directly proportional to the quantity of electricity passed through during electrolysis.

$$m \propto Q$$

$$m \propto It$$

Faraday second law of electrolysis: This law state that the number of Faraday required to discharge one mole of ion at an electrode equals to the number of charges on the ion

$$1\text{ mol of ion} = 1\text{ F}$$

$$2\text{ mol of ion} = 2\text{ F} \quad \text{Note: } 1\text{ F} = 96500\text{ C/mol}$$

$$\text{Recall, mole} = \frac{m}{M} \text{ therefore, } \frac{m}{M} = \frac{Q}{nF}$$

$$m = \frac{MQ}{nF} \quad \text{Where, } m = \text{mass (g)}, M = \text{Relative molecular mass of substance liberated}, n = \text{number of electron}, F = \text{faraday constant.}$$

PRACTICE QUESTIONS

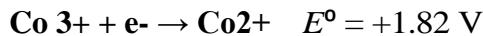
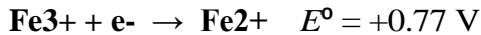
1. How long must a current of 2.50 A be applied to a solution of Ag^+ to plate out 11.5 g of Ag metal

2. What mass of Cu metal will plate out if a current of 4.50 A is applied to a solution of Cu(NO₃)₂ for a 35.0 min?
3. For the redox reaction, MnO₄⁻ + C₂O₄²⁻ + H⁺ → Mn²⁺ + CO₂ + H₂O

The correct coefficients of the reactants for the balanced equation are- (a) 16, 5, 2 (b) 2, 5, 16 (c) 2, 16, 5 (d) 5, 16, 2

Calculate the oxidation number of BOLD elements in the followings:

1. Na₂**B**₄O₇ 2. H₂S₂O₇ 3. H₄**P**₂O₇
2. In terms of electron transfer, what does an oxidizing agents do in redox reaction (a) donate electrons (b) accept electrons (c) donate a pair of electrons (d) accept a pair of electrons
3. The E^o (V) values for the two electrode are shown:



What is the EMF of the cell? (a) +2.59 V (b)+1.05 V (c) -1.05 V (d) -2.59 V

4. In voltaic cells, such as those diagrammed in your text, the salt bridge _____ (a) drives free electron from one half cell to another (b) is not necessary in other for the cell to work (c) act as a mechanism for the mixing of the two solutions (d) allow the charge balance to be maintained in the cell