Unit 12: Electrochemistry

Student Name: ____________________

Class Period: _3, 5, & 10_

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Unit 12 Vocabulary:

1. Alternating Current (AC): The current produced by power plants; the polarity (positive to negative current) shifts (altersates) at a household rate of 60 cycles per second (Hz).
2. Anode: The electrode at which oxidation occurs.
3. Cathode: The electrode at which reduction occurs.
4. Converter: A device that takes AC commercial current and converts it to DC current at the step-down voltage required by the device.
5. Direct Current (DC): The current produced by generators and batteries, where electricity flows only from anode to cathode. DC current is used in battery-powered electronics.
6. Electrolysis: The splitting apart of the elements in a compound by the application of electricity.
7. Electrolytic Cell: A process that uses electricity from an outside source to force a nonspontaneous redox reaction to occur. Examples of an electrolytic cell include recharging a battery, electrolytic decomposition of binary compounds, and electroplating of metals.
8. Electroplating: An electrolytic process that involves oxidizing a source metal into a solution using an external power source and then reducing the metal ion in solution onto a metallic object that you want plated in the source metal.
9. Half-reaction: A reaction that describes the change in oxidation number and the subsequent gain or loss of electrons that occurs during oxidation or reduction.
10. Load: A device or process that makes use of the electric current produced by an electrochemical cell.
11. Oxidation Number: The charge of an ion or the apparent charge of a nonmetal ion in a covalent bond.
12. Oxidizing Agent: The species that was reduced and therefore removes the electrons from the species that was oxidized.
13. Reducing Agent: The species that was oxidized and therefore gives electrons to the species that was reduced.
14. Salt Bridge: A semi-permeable barrier that allows the flow of ions from one half-cell to another half-cell, but prevents the direct mixing of the ions.
15. Species: The symbol and charge of an element or ion in a redox reaction.
16. Transformer: An electrical device that steps voltage up or down.
17. Voltaic Cell: An electrochemical cell that produces electrical current as a result of a spontaneous redox reaction. A voltaic cell is composed of two half-cells connected by a salt bridge, and two electrodes that connect to a load to complete the circuit.
## Unit 12 Homework Assignments:

<table>
<thead>
<tr>
<th>Assignment:</th>
<th>Date:</th>
<th>Due:</th>
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Notes page:
Electricity:

Why study electricity in chemistry? Isn’t that a physics topic? Well, yes it is, as I have taught Regents physics as well. But to understand what you can DO with electricity in physics, you need to understand how electricity is created in the chemical world.

In today’s (2015) world, electricity and electronics dominate our lives. I typed this unit packet on an electronic computer, looking at an electronic display, listening to Metallica that was recorded in a studio using electronic instruments. Barely a day goes by where anyone in this school will not interact SOMEHOW with something powered by electricity.

Before we dive into the depths of electrochemistry, let us take a look back at how elements gain and lose electrons, and the types of reactions we will be working with. Remember, these are REVIEW, not new materials.

If you need to look back, you can look at the following units:

- **Oxidation & Reduction**: Unit 5
- **Electronegativity & Ionization Energy**: Unit 5
- **Formation of Ions**: Unit 6
- **Trends in Ion size**: Unit 6
- **Types of Chemical Reactions**: Unit 8

By no means is this a comprehensive list of what to review; we started doing Regents review in early April for a reason.
## Electrochemistry Periodic Table Review

<table>
<thead>
<tr>
<th>Type of Element</th>
<th>Metal</th>
<th>Nonmetal</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electronegativity (and what it means for the element) - Table S</td>
<td>Metals have LOW electronegativity; they WEAKLY attract electrons in a chemical bond</td>
<td>Nonmetals have HIGH electronegativity; they STRONGLY attract electrons in a chemical bond</td>
</tr>
<tr>
<td>First Ionization Energy (and what it means for the element) - Table S</td>
<td>Metals have LOW first ionization energy; they can easily lose electrons when energy is added</td>
<td>Nonmetals have HIGH first ionization energies; they resist giving up valence electrons</td>
</tr>
<tr>
<td>How ions of this type of element are formed?</td>
<td>Metals LOSE valence electrons (oxidize) to lose a partially filled outer valence shell to form POSITIVE (less electrons than protons) CATIONS</td>
<td>Nonmetals GAIN valence electrons (reduction) to gain a stable octet of 8 valence electrons to form NEGATIVE (more electrons than protons) ANIONS</td>
</tr>
<tr>
<td>What happens to the atomic radius of this type of element as an ion is formed?</td>
<td>The RADIUS of the metal atom DECREASES as it loses electrons; the metal ION is SMALLER than the metal atom</td>
<td>The radius of the nonmetal atom increases as it gains electrons; the nonmetal ION is LARGER than the metal atom</td>
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</tbody>
</table>
## Redox Reaction Type Review

<table>
<thead>
<tr>
<th>Redox Rx Type</th>
<th>General Formula</th>
<th>Example</th>
<th>Uses in Electrochemistry</th>
</tr>
</thead>
<tbody>
<tr>
<td>Synthesis</td>
<td>A + B → AB</td>
<td>2 H₂ + O₂ → 2 H₂O</td>
<td>This reaction may be used to create electricity in a fuel cell. The e⁻ lost by H₂(g) will pass through a wire into the electric device, then the e⁻ pass into the gaining O₂(g)</td>
</tr>
<tr>
<td>Decomposition</td>
<td>AB → A + B</td>
<td>2 NaCl → 2 Na + Cl₂</td>
<td>This reaction is carried out by adding electricity to aqueous NaCl. The Na⁺⁺ is forced to gain e⁻ and reduce to Na⁰. The Cl⁻⁻ is forced to lose e⁻ and oxidize to form Cl₂⁰. This is electrolytic decomposition.</td>
</tr>
<tr>
<td>Single Replacement</td>
<td>A + BC → AC + B</td>
<td>Zn + Cu(NO₃)₂ → Zn(NO₃)₂ + Cu</td>
<td>This reaction is used to make voltaic cells (multiples of which when connected are called a “battery”). This example is called a zinc/copper cell, and it generates about 1.1 volts.</td>
</tr>
</tbody>
</table>
Redox Reactions in Electrochemistry:

- Redox reactions used for electrochemistry are driven by a change in charge among participating species.

<table>
<thead>
<tr>
<th>OXIDATION</th>
<th>Loss of electrons</th>
<th>Becomes more + charged</th>
</tr>
</thead>
<tbody>
<tr>
<td>REDUCTION</td>
<td>Gain of electrons</td>
<td>Becomes more - charged</td>
</tr>
</tbody>
</table>

Why do we need Oxidation numbers?

- The first step in determining which species was oxidized and which was reduced is knowing what the charge of each species starts at and ends at, so you can determine whether electrons were lost or gained. Why is that important? If you have ever put a battery in backwards in electronics, you’ll understand why. The device doesn’t work. Circuits are designed to only let electrons travel in ONE direction. Batteries are marked so we know which end is negative (the oxidizing anode) or which end is positive (the reducing cation). To do that, we need to know how to find the charges (oxidation numbers) of the active species involved.
Rules for Oxidation numbers:

1. The **SUM** of all charges in a **compound** equals **zero** (is neutral);
2. The **sum** of all charges in a **polyatomic** ion equals the charge of the polyatomic ion;
3. Uncombined (**single**, **diatomic**, or **network** solid) elements have a charge of **ZERO**;
   i. Na has a charge of zero;
   ii. O₂ has a charge of zero;
   iii. P₄ has a charge of zero;
4. **Fluorine** always has an oxidation number of **-1**;
5. **Oxygen** has an oxidation number of **-2** EXCEPT when in a peroxide, and then it is **-1**;
6. **Hydrogen** has an oxidation number of **+1** EXCEPT when combined with a metal forming a binary metal hydride;
7. If an **element** has **only** one charge listed on the periodic table that is the oxidation number for that element.
   i. Na has one charge listed of +1; Na has an oxidation number of +1;
   ii. Ca has one charge listed of +2; Ca has an oxidation number of +2;
   iii. Al has one charge listed of +3; Al has an oxidation number of +3.
8. If a nonmetal atom is the negative ion (anion) in a compound, then the top charge listed for that element on the periodic table is the oxidation number for that element.
   i. In NaCl, Cl is the negative ion (anion), and -1 is the top charge listed, so Cl has an oxidation number of -1;
   ii. In ZnS, S is the negative ion (anion), and -2 is the top charge listed, so S has an oxidation number of -2;
   iii. In Li$_3$N, N is the negative ion (anion), and -3 is the top charge listed, so N has an oxidation number of -3.

9. If an element in a compound of three or more elements has more than one charge listed, use the charges of the parts that you KNOW to determine the oxidation number of the unknown element.

Determining Oxidation Numbers using Rules:

I. Look at ZnSO$_4$: What is the oxidation number of sulfur in ZnSO$_4$?
   a. Zn has only one charge on the periodic table, +2, so Zn has an oxidation number of +2;
   b. (SO$_4^{-2}$) is a polyatomic ion, and the sum of all oxidation numbers in a polyatomic ion equals the charge of the polyatomic ion. The charge for (SO$_4^{-2}$) is -2, so ALL elements in (SO$_4^{-2}$) will add to a -2 oxidation number.
   c. Oxygen as stated will have an oxidation number of -2, so the four oxygen atoms will have a total of -8 oxidation number.
   d. As the (SO$_4^{-2}$) polyatomic ion needs a total -2 oxidation state, we now have a simple algebraic equation: -8 + X = -2
e. Solving for X, we get +6, so the oxidation number for sulfur in the polyatomic ion (SO$_4^{2-}$) is +6, and so sulfur will be +6 in ZnSO$_4$.

*Check the Periodic Table to see if sulfur has a +6 charge listed; it does!

II. Look at Fe(NO$_3$)$_2$: What are the oxidation numbers for Fe and N in Fe(NO$_3$)$_2$?

a. Look first at the nitrate (NO$_3^-$) polyatomic ion. The entire (NO$_3^-$) ion has a charge of -1, so the sum of all oxidation numbers in the (NO$_3^-$) polyatomic ion will be -1.

b. Again, oxygen has an oxidation number of -2. There are three oxygen atoms in the (NO$_3^-$) ion, for a total oxidation number of -6 for the oxygen.

c. Knowing we have -6 for the combined oxidation numbers of oxygen, and knowing that we need a -1 oxidation state for the entire (NO$_3^-$) ion, we again set up an algebraic equation: X + -6 = -1

d. Solving for X, we get +5, so the oxidation number for nitrogen in the polyatomic ion (NO$_3^-$) will be +5, and so nitrogen will be in Fe(NO$_3$)$_2$.

*Check the Periodic Table to see if nitrogen has a +5 charge listed; it does!

e. Now we can look at the oxidation number for iron. Any compound will have an overall zero (neutral) charge, so with two (NO$_3^-$) ions each having a charge of -1, the total negative oxidation number will be -2. This leaves Fe needing an oxidation number of +2 to zero out the charges on the compound.

*Check the Periodic Table to see if iron has a +2 charge listed; it does!
Practice Regents Questions-Oxidation Numbers (Ungraded):

1. In which compound does chlorine have the highest oxidation number?
   a) NaClO   b) NaClO₂   c) NaClO₃   d) NaClO₄

2. In which compound does carbon have an oxidation state of −4?
   a) CO   b) CO₂   c) CH₄   d) CCl₄

3. The oxidation number of nitrogen in N₂ is
   a) 0   b) +1   c) −3   d) +3

4. What is the sum of the oxidation number of atoms in the compound CO₂?
   a) 0   b) −2   c) −4   d) +4

5. In which compound is the oxidation number of oxygen treated as −1?
   a) CO   b) CO₂   c) H₂O   d) H₂O₂
Topic: **Identifying Redox Species**

**Objective:** How may we identify a species in a redox reaction?

**Identifying the Oxidizing and Reducing Species:**

- If the charge on a species becomes *more positive* going left to right, that species is *oxidized* and loses electrons;
- If the charge on a species becomes *more negative* going left to right, that species is *reduced* and gains electrons.

- For the reaction: \( \text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu(NO}_3\text{)}_2 + 2 \text{Ag} \)
  
  i. First determine the oxidation numbers for each species. If a polyatomic ion remains constant on both sides of the reaction, look up the charge for that polyatomic ion. In this case, nitrate (\(\text{NO}_3^-\)) is on both sides, so when we look on Table E we find the charge for the (\(\text{NO}_3^-\)) is -1.

  \[
  \text{Cu}^0 + 2 \text{Ag}^{+1}(\text{NO}_3^-)^{-1} \rightarrow \text{Cu}^{+2}(\text{NO}_3^-)_2^{-1} + 2 \text{Ag}^0
  \]

  ii. Note that the (\(\text{NO}_3^-\)) hasn’t changed charge.

  iii. The Cu goes from \(\text{Cu}^0\) to \(\text{Cu}^{+2}\), so \(\text{Cu}^0\) lost two \(e^-\) and is oxidized.

  iv. Each Ag goes from \(\text{Ag}^{+1}\) to \(\text{Ag}^0\), so each \(\text{Ag}^0\) gained one \(e^-\) and is reduced.
Redox Agents:

- Oxidizing Agent:
  i. The *species* in a reaction that *causes* the *oxidation* is known as the oxidizing agent. The oxidizing agent is the species that removes (accepts) electrons FROM the oxidizing species.

- Reducing Agent:
  i. The *species* in a reaction that *causes* the *reduction* is known as the reducing agent. The reducing agent is the species that donates (gives) electrons TO the reducing species.
• For the reaction: $\text{Cu}^0 + 2 \text{Ag}^+ (\text{NO}_3^-) \rightarrow \text{Cu}^{2+} (\text{NO}_3^-)_2 ^{-1} + 2 \text{Ag}^0$
  
  i. Cu$^0$ is oxidized, so Cu$^0$ is the reducing agent. Cu$^0$ causes Ag$^+$ to be reduced as Cu$^0$ gives electrons to the Ag$^+$.  
  
  ii. Ag$^+$ is reduced, so Ag$^+$ is the reducing agent. Ag$^+$ causes Cu$^0$ to be oxidized as Ag$^+$ takes electrons from the Cu$^0$.

• Spectator Ions:
  
  i. Note that the (NO$_3^-$) hasn’t changed in charge. The nitrate (in this reaction) is a *spectator* ion, as the NO$_3^-$ charge *remains* -1 on both sides.
Practice Regents Questions-Redox Species (Ungraded):

1. Which reaction is an example of an oxidation-reduction reaction?
   a) \( \text{AgNO}_3 + \text{KI} \rightarrow \text{AgI} + \text{KNO}_3 \)
   b) \( \text{Ba(OH)}_2 + 2 \text{HCl} \rightarrow \text{BaCl}_2 + 2 \text{H}_2\text{O} \)
   c) \( 2 \text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + \text{H}_2\text{O} \)
   d) \( \text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu(NO}_3)_2 + 2 \text{Ag} \)

2. Given the reaction: \( 2 \text{Al}_{(s)} + \text{Fe}_2\text{O}_3(s) + \text{heat} \rightarrow \text{Al}_2\text{O}_3(s) + 2 \text{Fe}_{(s)} \)
   Which species undergoes reduction?
   a) Al  
   b) Fe  
   c) \( \text{Al}^{+3} \)  
   d) \( \text{Fe}^{+3} \)

3. Given the reaction: \( \text{Zn}_{(s)} + 2 \text{HCl}_{(aq)} \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g) \)
   Which statement correctly describes what occurs when this reaction takes place in a closed system?
   a) There is a net loss of mass.
   b) There is a net gain of mass.
   c) Atoms of \( \text{Zn}_{(s)} \) gain electrons and are reduced.
   d) Atoms of \( \text{Zn}_{(s)} \) lose electrons and are oxidized.

4. Which equation represents an oxidation-reduction reaction?
   a) \( \text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O} \)
   b) \( 2 \text{HCl} + \text{FeS} \rightarrow \text{FeCl}_2 + \text{H}_2\text{S} \)
   c) \( 2 \text{HCl} + \text{CaCO}_3 \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2 \)
   d) \( 4 \text{HCl} + \text{MnO}_2 \rightarrow \text{MnCl}_2 + 2 \text{H}_2\text{O} + \text{Cl}_2 \)

5. Which reaction is an example of oxidation-reduction?
   a) \( 2 \text{KCl} \rightarrow 2 \text{K} + \text{Cl}_2 \)
   b) \( \text{KOH} + \text{HCl} \rightarrow \text{KCl} + \text{H}_2\text{O} \)
   c) \( \text{KCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{KNO}_3 \)
   d) \( \text{BaCl}_2 + \text{K}_2\text{SO}_4 \rightarrow 2 \text{KCl} + \text{BaSO}_4 \)
Topic: **Half-Reactions**

Objective: How might we break complete Redox reactions into steps?

**Half-Reactions:**

- A redox reaction may be split into two half-reactions, one half-reaction showing the oxidation process, and the other half-reaction showing the reduction process.

Watch Crash Course Chemistry Redox Reactions YouTube video

https://www.youtube.com/watch?v=lQ6FBA1HM3s&list=PL8dPuuaLiXtPHzzYuWy6fYEA9mQQ8oGr&index=10

1. For the reaction: \( \text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu(NO}_3)_2 + 2 \text{Ag} \)
   i. Write out the charges: \( \text{Cu}^0 + 2 \text{Ag}^{+1}(\text{NO}_3)^{-1} \rightarrow \text{Cu}^{+2}(\text{NO}_3)_2^{-1} + 2 \text{Ag}^0 \)
   ii. Oxidation Half-Reaction: \( \text{Cu}^0 \rightarrow \text{Cu}^{+2} + 2 e^- \)
   iii. This oxidation half-reaction shows that Cu goes from \( \text{Cu}^0 \) to \( \text{Cu}^{+2} \) by losing 2 electrons. The two electrons are placed on the products (right) side to a) show a loss, and b) show Conservation of Charge.
   iv. Reduction Half-Reaction: \( 2 \text{Ag}^{+1} + 2 e^- \rightarrow 2 \text{Ag}^0 \)
   v. This reduction half-reaction shows that each Ag goes from \( \text{Ag}^{+1} \) to \( \text{Ag}^0 \) by gaining 1 electron. The electrons are placed on the reactants (left) side to a) show a gain, and b) show Conservation of Charge.
   vi. Note that the spectator ion (NO\(_3\)) does NOT show in the half-reactions.
2. For the reaction: \( \text{F}_2 + \text{CaCl}_2 \rightarrow \text{CaF}_2 + \text{Cl}_2 \)
   i. Write out the charges: \( \text{F}_2^0 + \text{Ca}^{+2} \ 2 \text{Cl}^{-1} \rightarrow \text{Ca}^{+2} \text{F}_2^{-1} + \text{Cl}_2^0 \)
   ii. Oxidation Half-Reaction: \( 2 \text{Cl}^{-1} \rightarrow \text{Cl}_2^0 + 2 \text{e}^- \)
   iii. This oxidation half-reaction shows that each of the two chlorine ions loses an electron during the process of forming a neutral diatomic molecule of chlorine. The oxidized electrons are on the products side.
   iv. Reduction Half-Reaction: \( \text{F}_2^0 + 2 \text{e}^- \rightarrow 2 \text{F}^{-1} \)
   v. This reduction half-reaction shows that the neutral diatomic fluorine molecule gained a total of 2 electrons as it split to form two fluorine anions.

3. For the reaction: \( \text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3 \)
   i. Write out the charges: \( \text{N}_2^0 + 3 \text{H}_2^0 \rightarrow 2 (\text{N}^{-3} \ \text{H}_3^{+1}) \)
   ii. Oxidation Half-Reaction: \( 3 \text{H}_2^0 \rightarrow 3 \text{H}^{+1} + 6 \text{e}^- \)
   iii. Reduction Half-Reaction: \( \text{N}_2^0 + 6 \text{e}^- \rightarrow 2 \text{N}^{-3} \)
   iv. For the reaction: \( \text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \)
   v. Write out the charges: \( \text{Zn}^0 + 2 (\text{H}^{+1} \ \text{Cl}^{-1}) \rightarrow \text{Zn}^{+2} \ 2 \text{Cl}^{-1} + \text{H}_2^0 \)
   vi. Oxidation Half-Reaction: \( \text{Zn}^0 \rightarrow \text{Zn}^{+2} + 2 \text{e}^- \)
   vii. Reduction Half-Reaction: \( 2 \text{H}^{+1} + 2 \text{e}^- \rightarrow \text{H}_2^0 \)
   viii. \( \text{Cl}^{-1} \) has not changed charge; it is a spectator ion and need not be shown.
Balancing Redox Reactions:

- Previous examples have shown how the spectator ions in a redox reaction may be \textit{ignored} during redox reactions. We can therefore eliminate any \textit{spectator} ions from the balancing of half-reactions.
- For the reaction: \(2 \text{K} + \text{Ca(NO}_3\text{)}_2 \rightarrow 2 \text{KNO}_3 + \text{Ca}\)
  
  i. First, write out the charges:
  \[
  2 \text{K}^0 + \text{Ca}^{+2} (\text{NO}_3^-)_2 \rightarrow 2 \text{K}^{+1} \text{NO}_3^- + \text{Ca}^0
  \]

  ii. We can see that the nitrate (NO\(_3\)) is the spectator ion as it stays as a -1 charge: \(2 \text{K}^0 + \text{Ca}^{+2} (\text{NO}_3^-)_2 \rightarrow 2 \text{K}^{+1} \text{NO}_3^- + \text{Ca}^0\)

  iii. We can take the (NO\(_3\)) out of the net ionic reaction:
  \[
  2 \text{K}^0 + \text{Ca}^{+2} (\text{NO}_3^-)_2 \rightarrow 2 \text{K}^{+1} \text{NO}_3^- + \text{Ca}^0
  \]

  iv. That leaves us with: \(2 \text{K}^0 + \text{Ca}^{+2} \rightarrow 2 \text{K}^{+1} + \text{Ca}^0\)

  v. Notice that we are not only conserving mass (same number of K and Ca on both sides), but we are also conserving charge (total of +2 on each side).
Balancing Net Ionic Reactions:

I. When balancing net ionic reactions, as in any type of balancing, the numbers of a species **MUST** be equal on both sides of a reaction.

II. For net ionic reactions, the electrons are just another species introduced into the equation.

III. The number of electrons lost by the species being oxidized must equal the number of electrons gained by the species being reduced;

IV. The difference in charge between both sides of the oxidized species will be the coefficient of the reduced species;

V. The difference in charge between both sides of the reduced species will be the coefficient of the oxidized species.

1. Balance the net ionic redox reaction of: \( \text{Na}^0 + \text{Fe}^{+3} \rightarrow \text{Na}^{+1} + \text{Fe}^0 \)
   a) First, note that the charge on the reactants side is +3, while the charge on the products side is +1.
   b) Write the half-reactions:
      
      - Oxidation half-reaction: \( \text{Na}^0 \rightarrow \text{Na}^{+1} + 1 \text{ e}^- \)
      - Reduction half-reaction: \( \text{Fe}^{+3} + 3 \text{ e}^- \rightarrow \text{Fe}^0 \)
   c) Multiply the entire oxidation half-reaction by the coefficient of the free electrons in the reduction half-reaction:
\[ 3 \ (\text{Na}^0 \rightarrow \text{Na}^{+1} + 1 \ \text{e}^-) = 3 \ \text{Na}^0 \rightarrow 3 \ \text{Na}^{+1} + 3 \ \text{e}^- \]

d) Multiply the entire reduction half-reaction by the coefficient of the free electrons in the oxidation half-reaction:

\[ 1 \ (\text{Fe}^{+3} + 3 \ \text{e}^- \rightarrow \text{Fe}^0) = \text{Fe}^{+3} + 3 \ \text{e}^- \rightarrow \text{Fe}^0 \]

e) Place the reactants (left) side from BOTH equations together:

\[ (3 \ \text{Na}^0) + (\text{Fe}^{+3} + 3 \ \text{e}^-) \]

f) Place the products (right) side from BOTH equations together:

\[ (3 \ \text{Na}^{+1} + 3 \ \text{e}^-) + (\text{Fe}^0) \]

g) Place ALL reactants on the left side of the arrow, and ALL products on the right side of the arrow:

\[ (3 \ \text{Na}^0) + (\text{Fe}^{+3} + 3 \ \text{e}^-) \rightarrow (3 \ \text{Na}^{+1} + 3 \ \text{e}^-) + (\text{Fe}^0) \]

h) You can cancel out the free electrons (e-) from both sides:

\[ (3 \ \text{Na}^0) + (\text{Fe}^{+3} + 3 \ \text{e}^-) \rightarrow (3 \ \text{Na}^{+1} + 3 \ \text{e}^-) + (\text{Fe}^0) \]

i) That leaves you with the balanced NET ionic redox reaction:

\[ 3 \ \text{Na}^0 + \text{Fe}^{+3} \rightarrow 3 \ \text{Na}^{+1} + \text{Fe}^0 \]

(The charge on both sides is now +3)
2. Balance the net ionic redox reaction of: $\text{Au}^0 + \text{H}^{+1} \rightarrow \text{Au}^{+3} + \text{H}_2^0$

   a) First, note that the charge on the reactants side is +1, while the charge on the products side is +2.

   b) Write the half-reactions:

   Oxidation half-reaction: $\text{Au}^0 \rightarrow \text{Au}^{+3} + 3 \text{e}^-$

   Reduction half-reaction: $2 \text{H}^{+1} + 2 \text{e}^- \rightarrow \text{H}_2^0$

   c) Multiply the entire oxidation half-reaction by the coefficient of the free electrons in the reduction half-reaction:

   \[ 2 (\text{Au}^0 \rightarrow \text{Au}^{+3} + 3 \text{e}^-) = 2 \text{Au}^0 \rightarrow 2 \text{Au}^{+3} + 6 \text{e}^- \]

   d) Multiply the entire reduction half-reaction by the coefficient of the free electrons in the oxidation half-reaction:

   \[ 3 (2 \text{H}^{+1} + 2 \text{e}^- \rightarrow \text{H}_2^0) = 6 \text{H}^{+1} + 6 \text{e}^- \rightarrow 3 \text{H}_2^0 \]

   e) Place the reactants (left) side from BOTH equations together:

   \[ (2 \text{Au}^0) + (6 \text{H}^{+1} + 6 \text{e}^-) \]

   f) Place the products (right) side from BOTH equations together:

   \[ (2 \text{Au}^{+3} + 6 \text{e}^-) + (3 \text{H}_2^0) \]

   g) Place ALL reactants on the left side of the arrow, and ALL products on the right side of the arrow:

   \[ (2 \text{Au}^0) + (6 \text{H}^{+1} + 6 \text{e}^-) \rightarrow (2 \text{Au}^{+3} + 6 \text{e}^-) + (3 \text{H}_2^0) \]

   h) You can cancel out the free electrons ($\text{e}^-$) from both sides:

   \[ (2 \text{Au}^0) + (6 \text{H}^{+1} + 6 \text{e}^-) \rightarrow (2 \text{Au}^{+3} + 6 \text{e}^-) + (3 \text{H}_2^0) \]

   i) That leaves you with the balanced NET ionic redox reaction:

   \[ 2 \text{Au}^0 + 6 \text{H}^{+1} \rightarrow 2 \text{Au}^{+3} + 3 \text{H}_2^0 \]

   (The charge on both sides is now +6)
Practice Regents Questions-Redox Species (Ungraded):

1. Which half-reaction correctly represents reduction?
   a) $\text{Ag}^0 \rightarrow \text{Ag}^+ + e^-$
   b) $\text{F}_2 \rightarrow 2 \text{F}^- + 2 e^-$
   c) $\text{Fe}^{2+} + e^- \rightarrow \text{Fe}^{3+}$
   d) $\text{Au}^{3+} + 3 e^- \rightarrow \text{Au}^0$

2. Given the reaction: $\text{Fe}^0(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Fe}^{2+}(aq) + \text{Cu}^0(s)$
   Which half-reaction correctly shows that oxidation occurs?
   a) $\text{Fe}^0(s) \rightarrow \text{Fe}^{2+}(aq) + 2 e^-$
   b) $\text{Fe}^0(s) + 2 e^- \rightarrow \text{Fe}^{2+}(aq)$
   c) $\text{Cu}^{2+}(aq) + 2 e^- \rightarrow \text{Cu}^0(s)$
   d) $\text{Cu}^{2+}(aq) \rightarrow \text{Cu}^0(s) + 2 e^-$

3. Given the reaction: $2 \text{Al}^0(s) + 3 \text{Ni}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Ni}^0(s)$
   What is the total number of moles of electrons lost by two moles of $\text{Al}^0(s)$?
   a) 2
   b) 3
   c) 6
   d) 8

4. Which expression correctly represents a balanced reduction half-reaction?
   a) $\text{Na}^0 \rightarrow \text{Na}^+ + e^-$
   b) $\text{Na}^+ + e^- \rightarrow \text{Na}^0$
   c) $\text{Cl}_2 + 2 e^- \rightarrow \text{Cl}^-$
   d) $2 \text{Cl}^- \rightarrow \text{Cl}_2 + 2 e^-$

5. When an equation is correctly balanced, it must show conservation of
   a) Both charge and mass
   b) Mass but not of charge
   c) Charge but not of mass
   d) Neither charge nor mass
Oxidation Numbers and Half-Reactions Homework

Determine the oxidation numbers for each element in the following substances. Report each individual oxidation number for each element, not the total using subscripts. For these examples, oxygen (O) will be \(-2\); anything else is wrong.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Individual Ox #</th>
<th>Individual Ox #</th>
<th>Individual Ox #</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. N(_2)</td>
<td>N: 0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. PbO</td>
<td>Pb: +2</td>
<td>O: -2</td>
<td></td>
</tr>
<tr>
<td>4. Cu(ClO(_3))(_2)</td>
<td>Cu: +2</td>
<td>Cl: +5</td>
<td>O: -2</td>
</tr>
<tr>
<td>5. Cu(ClO(_2))(_2)</td>
<td>Cu: +2</td>
<td>Cl: +3</td>
<td>O: -2</td>
</tr>
<tr>
<td>6. Cu(ClO(_2))</td>
<td>Cu: +2</td>
<td>Cl: +1</td>
<td>O: -2</td>
</tr>
<tr>
<td>7. K</td>
<td>K: 0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>8. Pb(Cr(_2)O(_7))(_2)</td>
<td>Pb: +4</td>
<td>Cr: +6</td>
<td>O: -2</td>
</tr>
</tbody>
</table>

Balance the following redox reactions by the half-reaction method, re-writing the balanced half-equations with charges below the given unbalanced equation. Show your work for each reaction, placing the correct coefficients in the spaces provided.

1. \(1\) Cu + \(2\) Ag\(^{+1}\) \(\rightarrow\) \(1\) Cu\(^{+2}\) + \(2\) Ag

   \(Ox:\ Cu^0 \rightarrow 2\ e^- + Cu^{+2}\)

   \(Red:\ 2\ Ag^{+1} + 2\ e^- \rightarrow 2\ Ag^0\)

Cont’d on back
2. \( \underline{\text{2}} \) \( \underline{\text{2}} \) \( \underline{\text{3}} \) \( \underline{\text{3}} \) \( \underline{\text{Pb}^{2+}} \) \( \rightarrow \) \( \underline{\text{2}} \) \( \underline{\text{Fe}^{3+}} \) \( + \) \( \underline{\text{3}} \) \( \underline{\text{Pb}} \)

\textit{Ox:} \( \underline{\text{2 Fe}}^{0} \rightarrow \underline{\text{2 Fe}^{3+}} + \underline{6 e^{-}} \) \\
\textit{Red:} \( \underline{\text{3 Pb}^{2+}} + \underline{6 e^{-}} \rightarrow \underline{\text{3 Pb}^{0}} \)

3. \( \underline{\text{1}} \) \( \underline{\text{1}} \) \( \underline{\text{2}} \) \( \underline{\text{2}} \) \( \underline{\text{Li}}^{+2} \) \( \rightarrow \) \( \underline{\text{2}} \) \( \underline{\text{Li}^{+1}} \) \( + \) \( \underline{\text{1}} \) \( \underline{\text{2}} \) \( \underline{\text{Ni}}^{+2} \)

\textit{Ox:} \( \underline{\text{2 Li}^{0}} \rightarrow \underline{\text{2 Li}^{+1}} + \underline{2 e^{-}} \) \\
\textit{Red:} \( \underline{\text{Ni}^{+2}} + \underline{2 e^{-}} \rightarrow \underline{\text{Ni}^{0}} \)

Write ALL of the charges for each species above their symbols, then write the oxidation half-reaction and reduction half-reaction, identifying both the oxidizing and reducing agent, and any spectator ions (if present). 5 pts. ea.
**Topic: Electrochemistry & Potential**

**Objective:** How may we determine the potential for Redox reactions?

**Electrochemistry:**

The science of how to create voltaic cells that are combined into batteries is known as electrochemistry.

**Redox Potentials and Metal Activity:**

i. We’ve already used Reference Table J, *Activity Series*, when working with single replacement reactions. We learned that the higher species on Table J is more reactive than the lower species. This means that for metals, the *higher* up Table J the *greater* the potential to be *oxidized*. 
Topic: **Oxidation Potential**

Objective: How do we determine the redox potential for oxidation?

Which metal has the greatest potential to be oxidized?

- Lithium, Li, is the most active metal on Table J, and therefore Li has the greatest potential for losing electrons (oxidizing). Voltaic cells (batteries) made with lithium can produce a lot of voltage for their size. Lithium ion and lithium polymer batteries are also rechargeable.

- Which of the following reactions is more likely to occur?

1. \( \text{Cu}^{+2} + \text{Al}^{0} \rightarrow \text{Cu}^{0} + \text{Al}^{+3} \)  
   - Ox species: \( \text{Al}^{0} \)  
   - Red species: \( \text{Cu}^{+2} \)
   - i. Look on Table J; is Al higher than Cu? Yes, so Al is more reactive.
   - ii. As the more reactive species (Al) is being oxidized, this reaction works.

2. \( \text{Cu}^{0} + \text{Al}^{+3} \rightarrow \text{Cu}^{+2} + \text{Al}^{0} \)  
   - Ox species: \( \text{Cu}^{0} \)  
   - Red species: \( \text{Al}^{+3} \)
   - i. Look on Table J; is Cu higher than Al? No, Al is more reactive.
   - ii. The more reactive species (Al) would be reduced here; this is not likely.
   - iii. \( \text{Cu}^{+2} + \text{Al}^{0} \rightarrow \text{Cu}^{0} + \text{Al}^{+3} \) will be the reaction that occurs of these two.
Redox Potentials and Nonmetal Activity:

- We learned that the higher species on Table J is more reactive than the lower species in metals, and the same applies to nonmetals. This means that for nonmetals, the higher up Table J the greater the potential to be reduced.

Which nonmetal has the greatest potential to be reduced?

- F₂, the diatomic fluorine molecule, is the most active nonmetal on Table J, and therefore Li has the greatest potential for gaining electrons (reducing).

- Which of the following reactions is more likely to occur?

1. \( \text{Cl}_2^0 + 2 \text{F}^- \rightarrow 2 \text{Cl}^- + \text{F}_2^0 \)  
   Ox species: \( \text{F}_2^0 \)  
   Red species: \( \text{Cl}^- \)
   i. Look on Table J; is \( \text{F}_2 \) higher than \( \text{Cl}_2 \)? Yes, so \( \text{F}_2 \) is more reactive.  
   ii. The more reactive species (\( \text{F}_2 \)) would be oxidized here; this is not likely.

2. \( \text{F}_2^0 + 2 \text{Cl}^- \rightarrow 2 \text{F}^- + \text{Cl}_2^0 \)  
   Ox species: \( \text{Cl}^- \)  
   Red species: \( \text{F}_2^- \)
   i. Look on Table J; is \( \text{F}_2 \) higher than \( \text{Cl}_2 \)? Yes, \( \text{F}_2 \) is more reactive.  
   ii. As the more reactive species (\( \text{F}_2 \)) is reduced, this reaction works.  
   iii. \( \text{F}_2^0 + 2 \text{Cl}^- \rightarrow 2 \text{F}^- + \text{Cl}_2^0 \) will be the reaction that occurs of these two.
Practice Regents Questions-Activity Series Potential (Ungraded):

1. According to Reference Table J, which of these solid metals will react most readily with 1.0 M HCl\(_{(aq)}\) to produce H\(_2\) (g)?
   a) K  
   b) Ca  
   c) Mg  
   d) Zn

2. According to Reference Table J, which metal will react with Zn\(^{+2}\) but will not react with Mg\(^{+2}\)?
   a) Al\(_{(s)}\)  
   b) Ba\(_{(s)}\)  
   c) Cu\(_{(s)}\)  
   d) Ni\(_{(s)}\)

3. Under standard conditions, which pure metal will react with 0.1 M HCl to produce hydrogen gas?
   a) Ag  
   b) Au  
   c) Cu  
   d) Mg

4. Referring to Reference Table J, which reaction will not occur under standard conditions?
   a) Ba\(_{(s)}\) + 2 HCl\(_{(aq)}\) \rightarrow CaCl\(_2\)\(_{(aq)}\) + H\(_2\)(g)  
   b) Sn\(_{(s)}\) + 2 HCl\(_{(aq)}\) \rightarrow SnCl\(_2\)\(_{(aq)}\) + H\(_2\) (g)  
   c) Mg\(_{(s)}\) + 2 HCl\(_{(aq)}\) \rightarrow MgCl\(_2\)\(_{(aq)}\) + H\(_2\) (g)  
   d) Cu\(_{(s)}\) + 2 HCl\(_{(aq)}\) \rightarrow CuCl\(_2\)\(_{(aq)}\) + H\(_2\) (g)

5. According to Reference Table J, which reaction will take place spontaneously?
   a) Sr\(^{+2}\)\(_{(aq)}\) + Sn\(^0\)\(_{(s)}\) \rightarrow Sr\(^0\)\(_{(s)}\) + Sn\(^{+2}\)\(_{(aq)}\)  
   b) Ni\(^{+2}\)\(_{(aq)}\) + Pb\(^0\)\(_{(s)}\) \rightarrow Ni\(^0\)\(_{(s)}\) + Pb\(^{+2}\)\(_{(aq)}\)  
   c) Fe\(^{+2}\)\(_{(aq)}\) + Cu\(^0\)\(_{(s)}\) \rightarrow Fe\(^0\)\(_{(s)}\) + Cu\(^{+2}\)\(_{(aq)}\)  
   d) Au\(^{+3}\)\(_{(aq)}\) + Al\(^0\)\(_{(s)}\) \rightarrow Au\(^0\)\(_{(s)}\) + Al\(^{+3}\)\(_{(aq)}\)
Voltaic Cells:

- A **Voltaic** Cell is composed of *two* separate redox *half-reactions*, with the electrons from the oxidizing half-reaction powering a load, then returning to the reduction half-reaction.
- Multiple voltaic cells comprise a chemical battery.

Constructing a Voltaic Cell:

i. Each *half-reaction* is carried out in a *half-cell*. In these half-cells, the uncombined free element acts as the electrode which is immersed in a solution containing the ion of that element.

ii. The half-cells are connected at their electrodes by a *wire* that is used to *transport* the *electrons* from the oxidized electrode (anode) to the reduced electrode (cathode).

iii. In the circuit between the anode and cathode is a *load*, the device that actually *uses* the *current* produced by the difference between the oxidation potential and the reduction potential of the two electrodes.

iv. Connecting the half-cells directly is a *salt bridge* that *completes* the *circuit*. The salt bridge is made of porous material that contains a salt with a different cation than the cations in the electrodes.
Labeling an Electrochemical Cell:

i. The **Ox**idation Half-Cell is called the **Anode** (An Ox);

ii. The **Red**uction Half-Cell is called the **Cat**hode (Red Cat);

iii. Electrons travel from the anode (negative), through the load, and to the cathode (positive);

iv. Anions travel across the salt bridge from cathode to anode.
Diagram of Daniell Cell: $\text{Zn} + \text{Cu(NO}_3\text{)}_2 \rightarrow \text{Zn(NO}_3\text{)}_2 + \text{Cu}$

Typical Wet Cell Schematic

\begin{align*}
\text{Anode Reaction:} & \quad \text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2e^- \\
\text{Cathode Reaction:} & \quad \text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu(s)}
\end{align*}
Voltaic Cell Function:

What is going on inside a Voltaic Cell? (Previous page diagrams)

- The **OVERALL** process may be summed up as this: The Zn\textsuperscript{0} metal, which is the *anode* (oxidizing electrode, or negative terminal of the cell), *loses* electrons (\(e^-\)), which the Cu\textsuperscript{+2}\textsubscript{(aq)} ions in the reduction half-cell are “pulling” through the wire.

- Steps in the Voltaic Cell Process:
  1. The Zn\textsuperscript{0} metal forms Zn\textsuperscript{+2}\textsubscript{(aq)}, and dissolves into aqueous solution (Zn\textsuperscript{+2}\textsubscript{(aq)}, and the zinc anode loses material over time;
  2. The electrons from the zinc anode pass through the metal wire, through the load (voltmeter or other device), and into the copper cathode;
  3. The Cu\textsuperscript{0} cathode has no use for the \(e^-\) from the other half-cell, so the \(e^-\) line the surface of the copper cathode;
  4. The Cu\textsuperscript{+2}\textsubscript{(aq)} ions in the reduction half-cell solution are attracted to the \(e^-\) on the surface of the cathode, and when the Cu\textsuperscript{+2}\textsubscript{(aq)} ions touch the surface of the copper cathode, the jump to the cations, reducing Cu\textsuperscript{+2}\textsubscript{(aq)} to Cu\textsuperscript{0};
  5. The newly reduced Cu\textsuperscript{0} atoms attach to the surface of the copper cathode, increasing the mass of the copper cathode;
6. The nitrate \((\text{NO}_3^-)_{\text{aq}}\) ions in solution travel across the salt bridge from the copper cathode to the zinc anode. More \(\text{Zn}^{+2}_{\text{aq}}\) ions in the oxidation half-cell requires more nitrate ions to cancel out the charges. The concentration of \(\text{Zn(NO}_3\text{)}_2\text{(_{aq})}\) in the oxidation half-cell increases; 

7. The concentration of \(\text{Cu(NO}_3\text{)}_2\text{(_{aq})}\) in the reduction half-cell decreases. As the concentration of \(\text{Cu(NO}_3\text{)}_2\text{(_{aq})}\) and the mass of \(\text{Zn}^0\) decrease, the potential difference decreases as well. When either \(\text{Zn}^0\) or \(\text{Cu(NO}_3\text{)}_2\text{(_{aq})}\) are gone, the voltaic cell is ‘dead’. 

8. A ‘dead’ voltaic cell occurs when the concentration of ions in both cells has reached equilibrium, and the potential voltage has then reached zero. The ‘dead’ voltaic cell may be ‘recharged’ by forcing the \(e^-\) from the cathode and back to the \(\text{Zn}^{+2}_{\text{aq}}\). This is done with a recharger that reverses the electron flow.
Topic: **Designing a Wet Voltaic Cell**

Objective: How do we decide WHAT to make a Voltaic Cell from?

Designing a Wet Voltaic Cell:

- When designing a wet *voltaic* cell, you need to *choose* an *anode*, a *cathode*, and two *aqueous* ionic *solutions* that each contains the same element as the anode you have chosen AND the same spectator ion.

1. Choose the Metal Electrodes:
   i. More active (oxidizing-Table J) metal for negative anode: Aluminum (Al\(^0\))
   ii. Less active (reducing-Table J) metal for positive cathode: Silver (Ag\(^0\))

2. Choose the solutions for each half-cell:
   i. Al\(^0\) has a \(^+3\) oxidation number; use three (NO\(_3^-\)) to make Al(NO\(_3\))\(_3(aq)\) solution that goes into the anode (oxidizing) half-cell
   ii. Ag\(^0\) has a \(^+1\) oxidation number; use one (NO\(_3^-\)) to make AgNO\(_3(aq)\) solution that goes into the cathode (reducing) half-cell
Voltaic Reaction: $\text{Al} + 3 \text{AgNO}_3 \rightarrow \text{Al(NO}_3)_3 + 3 \text{Ag}$

Write out the charges: $\text{Al}^0 + 3 (\text{Ag}^{+1} \text{NO}_3^{-1}) \rightarrow (\text{Al}^{+3} 3 \times \text{NO}_3^{-1}) + 3 \text{Ag}^0$

- The oxidized metal ($\text{Al}^0$, anode) becomes smaller as it forms aqueous ions. The reduced metal ($\text{Ag}^0$, cathode) becomes larger as it gathers reduced metal ions from solution.
- The negative ($\text{NO}_3^-$) anions pass through the salt bridge from cathode ($\text{Ag}^0$ half-cell) to anode ($\text{Al}^0$ half-cell).
• The positive (Al\(^{+3}\)) cations pass through the salt bridge from anode (Al\(^0\) half-cell) to cathode (Ag\(^0\) half-cell)

• The salt bridge may contain any salt solution (EXCEPT one containing Al\(^{+3}\) or Ag\(^{+1}\)). If we use NaCl\(_{aq}\) in this voltaic cell, when the circuit is FIRST connected, Na\(^{+1}\) passes into the cathode cell, and Cl\(^{-1}\) passes into the anode cell. Once the charge ‘flow’ is established, the NO\(_3^-\) anions and Al\(^{+3}\) cations may travel the salt bridge.
Practice Regents Questions—Voltaic Cells (Ungraded):

1. What condition is indicated when a chemical voltaic cell’s voltage has dropped to zero?
   a) *The cell reaction has reached equilibrium.*
   b) The cell reaction has completely stopped.
   c) The concentration of the reactants has increased.
   d) The concentration of the products has decreased.

2. Which statement is true for any electrochemical cell?
   a) Reduction occurs at the anode, only.
   b) *Oxidation occurs at the anode, only.* An Ox
   c) Oxidation occurs at both the anode and the cathode.
   d) Reduction occurs at both the anode and the cathode.

Base your answers to questions 3 and 4 below on the equation and diagram representing an electrochemical cell at 298 K and 1 atm.

3. Which species is oxidized when the switch is closed?
   a) $\text{Ag}^0_{(s)}$
   b) $\text{Mg}^0_{(s)}$
   c) $\text{Ag}^{+1}_{(aq)}$
   d) $\text{Mg}^{+2}_{(aq)}$

4. When the switch is closed, electrons flow from
   a) $\text{Ag}^0_{(s)}$ to $\text{Mg}^0_{(s)}$
   b) $\text{Mg}^0_{(s)}$ to $\text{Ag}^0_{(s)}$
   c) $\text{Ag}^{+}_{(aq)}$ to $\text{Mg}^{+2}_{(aq)}$
   d) $\text{Mg}^{+2}_{(aq)}$ to $\text{Ag}^+_{(aq)}$
Electrochemistry and Voltaic Cells Homework

Write the symbol and charge of a Group 1 metal ion that has a greater potential to undergo oxidation than sodium: \( \text{Li}^+, \text{Rb}^+, \text{K}^+, \text{Cs}^+ \)

Write the symbol and charge of a nonmetal ion that has a greater potential to undergo oxidation than chlorine: \( \text{Br}^-, \text{I}^- \)

Using the table below, when the following pairs of elements undergo a reaction, indicate which element will be oxidized and which element will be reduced according to Reference Table J. (1 pt. ea.)

<table>
<thead>
<tr>
<th>Element pair</th>
<th>Oxidized</th>
<th>Reduced</th>
<th>Element pair</th>
<th>Oxidized</th>
<th>Reduced</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Mg}^0 ) and ( \text{Cu}^0 )</td>
<td>( \text{Mg} )</td>
<td>( \text{Cu} )</td>
<td>( \text{Ni}^0 ) and ( \text{Pb}^0 )</td>
<td>( \text{Ni} )</td>
<td>( \text{Pb} )</td>
</tr>
<tr>
<td>( \text{Zn}^0 ) and ( \text{Li}^0 )</td>
<td>( \text{Li} )</td>
<td>( \text{Zn} )</td>
<td>( \text{Na}^0 ) and ( \text{Cu}^0 )</td>
<td>( \text{Na} )</td>
<td>( \text{Cu} )</td>
</tr>
<tr>
<td>( \text{Na}^0 ) and ( \text{Ca}^0 )</td>
<td>( \text{Ca} )</td>
<td>( \text{Na} )</td>
<td>( \text{Cr}^0 ) and ( \text{Sr}^0 )</td>
<td>( \text{Sr} )</td>
<td>( \text{Cr} )</td>
</tr>
<tr>
<td>( \text{Mn}^0 ) and ( \text{Ba}^0 )</td>
<td>( \text{Ba} )</td>
<td>( \text{Mn} )</td>
<td>( \text{Au}^0 ) and ( \text{Ag}^0 )</td>
<td>( \text{Ag} )</td>
<td>( \text{Au} )</td>
</tr>
</tbody>
</table>

On the nickel/barium voltaic cell diagram on the next page, label each:

- Anode, cathode, + electrode, - electrode, electron flow direction, anion flow direction, salt bridge, and load

Use Reference Table J to determine your choices.

1. According to Table J, \( \text{Ba}^0 \) will undergo oxidation and \( \text{Ni}^{+2} \) will undergo reduction.
2. Label the diagram below according to the directions on the previous page.

3. Oxidation Half-Reaction: \( \text{Ba}^0 \rightarrow \text{Ba}^{2+} + 2 \text{e}^- \)

4. Reduction Half-Reaction: \( \text{Ni}^{2+} + 2 \text{e}^- \rightarrow \text{Ni}^0 \)

5. Oxidizing Agent: \( \text{Ni}^{2+} \)

6. Reducing Agent: \( \text{Ba}^0 \)

7. In which direction are the cations crossing the salt bridge? \( \text{anode} \rightarrow \text{cathode} \)

8. Could \( \text{Ba(NO}_3\text{)}_2(\text{aq}) \) be used in the salt bridge for this cell? Why, or why not? 
   
   \text{No, you need a soluble salt that contains different metals than either the anode or cathode}

9. Which electrode would become \textit{smaller} during the operation of this cell?
   
   \text{Barium anode}

10. Which electrode would become \textit{larger} during the operation of this cell?

    \text{Nickel cathode}
**Using Electricity to Decompose Compounds:**

- Many elements that we want to work with are only found in nature in compounds. We can use *electricity* to *separate* these *elements*.

<table>
<thead>
<tr>
<th>Group #</th>
<th>Group Name</th>
<th>Group element properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Alkali metals</td>
<td>- Form $+1$ ions by loss of only valence $\text{e}^-$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- Have the lowest electronegativity and lowest first ionization energy values on period table</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- Have the largest atomic radii on period table</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- React violently with water producing strong bases and $\text{H}_2(\text{g})$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- Are ONLY found in compounds in nature</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- May be extracted from compounds using electrolytic reduction</td>
</tr>
<tr>
<td>2</td>
<td>Alkaline Earth metals</td>
<td>- Form $+2$ metals by lose of both valence $\text{e}^-$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- Have the 2$^{\text{nd}}$ lowest electronegativity and 2$^{\text{nd}}$ lowest first ionization energy values on periodic table</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- React quickly with water to produce a weak base and $\text{H}_2(\text{g})$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- Are ONLY found in compounds in nature</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- May be extracted from compounds using electrolytic reduction</td>
</tr>
<tr>
<td>3-12</td>
<td>Transition metals</td>
<td>- Can lose from valence shell and the outermost kernel shell to form more than one positive charge</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- Form colored compounds</td>
</tr>
<tr>
<td></td>
<td></td>
<td>- May be removed from compounds using complex reactions, including single replacement</td>
</tr>
<tr>
<td>Group #</td>
<td>Group Name</td>
<td>Group element properties</td>
</tr>
<tr>
<td>--------</td>
<td>---------------</td>
<td>------------------------------------------------------------------------------------------</td>
</tr>
<tr>
<td>17</td>
<td>Halogens</td>
<td>• Form -1 ions by gaining one e⁻ to create a stable valence octet</td>
</tr>
<tr>
<td></td>
<td></td>
<td>• Have the highest electronegativity and 2&lt;sup&gt;nd&lt;/sup&gt; highest first ionization energy on the period table</td>
</tr>
<tr>
<td></td>
<td></td>
<td>• Have the smallest atomic radii on the periodic table</td>
</tr>
<tr>
<td></td>
<td></td>
<td>• React violently with metals to form a salt</td>
</tr>
<tr>
<td></td>
<td></td>
<td>• Are ONLY found in compounds in nature</td>
</tr>
<tr>
<td></td>
<td></td>
<td>• May be extracted from compounds using electrolytic oxidation</td>
</tr>
<tr>
<td>18</td>
<td>Noble Gases</td>
<td>• Do not readily form ions, as they have a stable valence octet</td>
</tr>
<tr>
<td></td>
<td></td>
<td>• Have the highest first ionization energy on the periodic table</td>
</tr>
<tr>
<td></td>
<td></td>
<td>• Are found in pure monoatomic molecules in nature</td>
</tr>
</tbody>
</table>

- Even though Group 1 metals, Group 2 metals, and Group 17 nonmetals are only found in compounds in nature, pure forms of these elements have many uses.
- Extracting these highly reactive elements may be done by electrolysis.
- Electrolysis involves melting the solid ionic compound containing the element you want to extract and then passing a source of direct current through the liquid. Remember that ionic solids melt at HIGH temperatures; we could not melt ionic solids in the lab using Bunsen burners, so this is a high-temperature process.
- The current must be DC, as the current flow needs to be one-way, to allow the elements to be separated.
Electrolysis of pure NaCl\(_{(l)}\):

- Note that this example shows the electrolytic decomposition of NaCl, but the same process may be used for the extraction of any Group 1, 2 or 17 element from a compound.

- Note that this is NOT aqueous NaCl\(_{(aq)}\); the solid NaCl\(_{(s)}\) has been heated (above 1074 K) and is molten NaCl\(_{(l)}\).

- This is the industrial process used to extract sodium and chlorine from molten sodium chloride. The solid sodium chloride is heated in a furnace to the molten state (1074 K for NaCl), then high-temperature metal electrodes are placed in the molten liquid.

- Why use electrolysis to force the decomposition of sodium chloride? Nature tends toward lower energy states (decreasing enthalpy).
i. Group 1, 2, and 17 elements are highly unstable when alone. If they form a compound with another element, they decrease energy (lower enthalpy).

ii. By separating the elements with electricity, we force them apart (add potential energy; increasing enthalpy).

iii. This higher-energy state of the pure elements means that they will readily interact with ANY material they can to form a compound and shed the extra energy. Remember the lithium and sodium we put in water? It burned in WATER! How about the magnesium ribbon after using just a candle? We got a LOT of energy (heat & light) out of that! (Some of you STILL complain of being blinded!)

**Reaction of NaCl Electrolysis:**

- The negative electrode from the direct-current (DC) supply provides the electrons that reduce the positive metal cation to form a pure metal, using NaCl\(_{(s)}\) as an example:
  - Na\(^{+1}\)\(_{(i)}\) + 1 e\(^-\) (from DC supply) \(\rightarrow\) Na\(^0\)\(_{(s)}\) (pure sodium metal)

- The positive electrode from the DC supply removes electrons from the negative nonmetal anion, with the nonmetal anion being oxidized to form a pure nonmetal diatomic gas molecule:
• 2 Cl\textsuperscript{-1(\text{aq})} \rightarrow \text{Cl}_2\textsuperscript{0(\text{g})} + 2 \text{e}^{-\text{(aq)}} (\text{these electrons are pulled to the positive DC supply electrode})

• In this example, pure NaCl\textsubscript{(s)} was used, but the process will work for any binary compound of an alkali metal and halogen. The metal will reduce on the cathode, and the nonmetal will oxidize on the anode. As long as a continuous input of DC current is applied to a molten mixture, you will form a halogen gas (\text{Cl}_2\textsuperscript{0(\text{g})} in this example) and a pure metal (\text{Na}\textsuperscript{0(\text{s})} in this example).

• This is a nonspontaneous reaction as it stops as soon as the energy is shut down.
Topic: **Electrolysis of Water**

Objective: What process is used to split water into its components?

Electrolysis of liquid water (H$_2$O$_{(l)}$)

- Water is comprised of the most abundant element in the known Universe, hydrogen, and the most abundant element on planet Earth, oxygen. Wherever we go, we find water in some form. This is good, as we require it for life.

- What other uses does water have? We can’t burn it. Or can we? Not directly, but if we *separate* water *into* its two components, *hydrogen* gas (H$_2$(g)) and *oxygen* gas (O$_2$(g)), both of those gases are reactive and can be used for generating power, either by combustion, or other means.
Electrons come from the negative pole of the DC supply and travel into the cathode in the aqueous sulfuric acid solution. The dilute \( \text{H}_2\text{SO}_4(aq) \) acts as an electrolyte only; pure water won’t conduct electricity. The \( e^- \) traveling into the cathode attracts the positively charge \( \text{H}^+ \) cations. The \( e^- \) force the \( \text{H}^+ \) cations to reduce forming diatomic hydrogen gas, which bubbles through the solution and is collected. The reduction half-reaction is:

\[
2 \text{H}^+ + 2 e^- \rightarrow \text{H}_2^0
\]

Note in the diagram above that there is TWICE as much space above the negative electrode (cathode). That is due to the fact that since the molecular formula for water is \( \text{H}_2\text{O} \), for each mole of water decomposed by electrolysis, we end up with two moles of hydrogen gas and one mole of oxygen gas.

\[
2 \text{H}_2\text{O}(d) + \text{DC} \rightarrow 2 \text{H}_2(g) + \text{O}_2(g)
\]

Electrons are also being pulled towards the positive pole of the DC supply. The \( \text{O}^-2 \) ions are forced to give up their extra \( e^- \) and be oxidized forming diatomic oxygen gas, which bubbles through the solution and is collected. The pulled from the \( \text{O}^-2 \) travel back into the positive pole of the DC supply, completing the circuit. The oxidation half-reaction is:

\[
2 \text{O}^-2 \rightarrow \text{O}_2^0 + 4 e^-
\]
Practice Regents Questions - Voltaic Cells (Ungraded):

1. The reaction $2 \text{H}_2\text{O}(l) \rightarrow 2 \text{H}_2(g) + \text{O}_2(g)$ is forced to occur by use of an externally applied electric current. This process is named
   a) Hydrolysis  
   b) Electrolysis  
   c) Precipitation  
   d) Neutralization

2. Which net reaction occurs by the process of electrolysis?
   a) $2 \text{H}_2\text{O}(l) \rightarrow 2 \text{H}_2(g) + \text{O}_2(g)$
   b) $2 \text{HgO}_2(s) \rightarrow 2 \text{HgO}(l) + \text{O}_2(g)$
   c) $\text{MgCO}_3(s) \rightarrow \text{MgO(s)} + \text{CO}_2(g)$
   d) $2 \text{KClO}_3(l) \rightarrow 2 \text{KCl}(s) + 3 \text{O}_2(g)$

3. Metals from which groups listed below may be obtained by the reduction of their fused salts?
   a) Group 1 and Group 2
   b) Group 2 and Group 11
   c) Group 1 and Group 12
   d) Group 11 and Group 12

4. Which element listed below is obtained only by the electrolysis of its fused salt?
   a) Zinc  
   b) Gold  
   c) Silver  
   d) Lithium

5. Which equation represents the half-cell reaction that occurs at the negative electrode during the electrolysis of calcium chloride?
   a) $2 \text{Cl}^-(aq) \rightarrow \text{Cl}_2(g) + 2 \text{e}^-(aq)$
   b) $\text{Ca}^{+2}(aq) \rightarrow \text{Ca}^0(s) + 2 \text{e}^-(aq)$
   c) $\text{Ca}^{+2}(aq) + 2 \text{e}^-(aq) \rightarrow \text{Ca}^0(s)$
   d) $2 \text{Cl}^-(aq) \rightarrow \text{Cl}_2(g) + 2 \text{e}^-(aq)$
**Electroplating of Metals**

Objective: How can we bond one metal over a different metal?

**Electroplating:**

- Metals are really cool materials to work with. We can bend them, turn them into wires, melt and cast them, and form them into very thin sheets. Still they have issues with the environment around them, in that almost all metals want to give electrons up (oxidize) to the environment. One way we might limit oxidation of a metal is to **electroplate** it with a **layer** of less reactive metal over the cheaper, but stronger metal underneath.

- Examples of metals that are added to cheaper, more durable metals are chromium and gold. Both of these metals are highly resistant to oxidation, which is why many cars, motorcycles, and bathroom fixtures have electroplated metal (yes, bathroom fixtures may be gold plated; I was in one of Saddam Hussein’s summer palaces in Iraq, and the bathroom had gold-plated fixtures!)
Electroplating of one metal over another metal:

**Electroplating (Electrolytic) Reaction Process:**

*Note: Be AWARE that the Anode and Cathode have **REVERSED** polarity in *electroplating* from their roles in a wet voltaic cell.*

This makes sense; instead of generating electricity by breaking down (electrolytes and electrodes), we are using electricity to add new material (electrolytic process).

1. Silver is being oxidized at the positive electrode (anode) as the e\(^-\) are stripped off and pulled towards the positive end of the DC supply (up the wire). The silver enters the silver nitrate (AgNO\(_3\)(aq)) solution as additional silver cations (Ag\(^+\)(aq)):

\[
Ag^0_{(s)} \rightarrow Ag^{+1}_{(aq)} + 1 e^{-}_{(aq)}
\]
2. At the negative electrode (cathode) electrons are coming from the DC supply and being forced into the object being plated. The shared nature of the electrons in a metallic bond does not allow the metal atoms to gain electrons; these extra electrons go to the surface of the metal. Once at the metal object’s surface, any Ag\(^{+1}\) ion that touches the surface (and the extra e\(^-\)) are reduced to Ag\(^0\) metal, and adhere to the surface of the object being plated:

\[ \text{Ag}^{+1}_{(aq)} + \text{e}^-_{(aq)} \rightarrow \text{Ag}^0_{(s)} \]

3. The longer the current is flowing, and as long as there is still silver in the anode, the plating process will continue.

Electroplating Decomposition examples:

1. What solution could you use to copper-plate a zinc kitchen fixture?
   i. To copper-plate zinc, you require an aqueous solution containing a salt of the plating metal. To plate copper, choose a soluble salt (Reference Table F); copper (II) nitrate would work well.

2. To chromium-plate an aluminum trim piece, how would you hook up the electrodes to the DC power supply?
   i. The object to be plated becomes the cathode, which is attached to the negative end of the DC supply. The chromium metal would become the anode, which is attached to the positive end of the DC supply. In a solution containing a chromium salt, the chromium metal ions will be reduced onto the object to be plated.

Watch Bozeman Chemistry Electrochemistry YouTube video

https://www.youtube.com/watch?v=Rt7-VrmZuds&list=PLIIVwaZQkS2op2kDuFifhStNsS49LAxkZ&index=34
Practice Regents Questions-Voltaic Cells (Ungraded):

1. Which statement describes the redox reaction that occurs when an object is electroplated?
   a) It is spontaneous and requires an electric current.
   b) It is spontaneous and produces an electric current.
   c) It is non-spontaneous and requires an electric current.
   d) It is non-spontaneous and produces an electric current.

2. A metal object is to be electroplated with silver. Which set of electrodes should be used?
   a) A silver anode and a metal object as the cathode
   b) A silver cathode and a metal object as the anode
   c) A platinum anode and a metal object as the anode
   d) A platinum cathode and a metal object as the anode

3. What occurs when an electrolytic cell is used for silver-plating a spoon?
   a) A reduction reaction takes place at the anode.
   b) A chemical reaction produces an electric current.
   c) An oxidation reaction takes place at the cathode.
   d) An electric current produces a chemical reaction.
4. Which statement best describes the key in the diagram below?

![Diagram of an electrochemical cell]

a) It acts as the anode and is positive.
b) It acts as the anode and is negative.
c) It acts as the cathode and is positive.
**d) It acts as the cathode and is negative.**

5. In an electrolytic cell, a negative ion will migrate to and undergo oxidation at the
a) Anode, which is negatively charged
**b) Anode, which is positively charged**
c) Cathode, which is positively charged
d) Cathode, which is negatively charged
Notes page:
Decomposition (Electrolytic) and Electroplating Homework

Complete and balance the decomposition reaction of potassium bromide:

\[ \underline{2} \text{KBr} \rightarrow \underline{2} \text{K} + \underline{1} \text{Br}_2 \]

Write the oxidation and reduction Half-Reactions that occur when potassium bromide is decomposed. Label the diagram, with arrows for e⁻ flow direction.

**Oxidation Half-Reaction:**

\[ \underline{2} \text{Br}^- \rightarrow \text{Br}_2^0 + 2 \text{e}^- \]

**Reduction Half-Reaction:**

\[ \underline{2} \text{K}^+ + 2 \text{e}^- \rightarrow \underline{2} \text{K}^0 \]

_\text{K}^0 \text{metal}_ \text{ forms here at the cathode, which is the (-) electrode}_

_\text{Br}_2^0 \text{ gas}_ \text{ forms here at the anode, which is the (+) electrode}_

Cont'd next page
In the diagram below, copper is being plated onto a zinc penny.

1. Complete the partial words or phrases in the provided spaces.
2. Draw arrows showing the e- flow.
3. Choose a suitable soluble compound for the aqueous solution.