Unit 8
Redox and Electrochemistry

Text Chapter 19 & 20
Reference Table J
Oxidation

- Involves a loss of electrons or any chemical change in which an oxidation number increases that is, becomes more positive or less negative.

- Example: Changes from -3 to -1, -1 to +2, +2 to +7.
Reduction

- Involves a gain of electrons or any chemical change in which an oxidation number decreases that is, becomes more negative or less positive.

- Example: Changes from +1 to -1, +2 to -1, +7 to +2.
Mnemonic Device to help you remember!

- **LEO GER** –
  - **L**ose **E**lectrons **O**xidation
  - **G**ain **E**lectrons **R**eduction
Redox

- Oxidation and Reduction Occur Together
Oxidation Numbers

- the charge assigned to an atom in a particular substance that represents the number of electrons that must be added to or removed from an atom in a combined state to convert the atom into the elemental form.
Rules for Assigning Oxidation Numbers

1. the oxidation number of an atom of any free (uncombined) state is zero (Examples: Fe, Cu, H₂O₂F₂Br₂I₂N₂Cl₂)
2. The oxidation number of all group 1 elements is +1.
3. The oxidation number of all group 2 elements is +2
4. Fluorine always has an oxidation number of −1.
5. Oxygen usually has an oxidation number of −2 except with: 1) Fluorine it has a +2 (ex. OF₂) and 2) in peroxides it has a −1 (H₂O₂)
6. Hydrogen usually has a +1 oxidation number Except with metal hydrides (Group 1 & 2 metals with H)

*Use these rules DO NOT rely on your Reference Tables as only “Select Oxidation states” are listed NOT ALL of them!
The sum of the oxidation numbers in a compound must equal zero.

Examples of Oxidation numbers in compounds:

\[ \text{Na}_2\text{S} \]
Oxidation number of Na: +1 → Total charge of all Na’s: +2
Oxidation number of S: -2 → Total charge of all S’s: -2
Total charge of compound = 0

\[ \text{H}_2\text{SO}_4 \]
Oxidation number of H: +1 → Total charge of all H’s (2 x+1): +2
Oxidation number of O: -2 → Total charge of all O’s (4 x-2): -8
Oxidation number of S: +6 → Total charge of all S’s: +6
Total charge of compound = 0
The sum of the oxidation numbers in an ion must equal the charge on the ion.

\[ \text{SO}_3^{2-} \]

Oxidation number of O: \(-2\) → Total charge of all O’s: \((3 \times -2)\) \(-6\)

Oxidation number of S: \(+4\) → Total charge of all S’s: \(+4\)

Total charge of ion = \(-2\)

\[ \text{PO}_4^{3-} \]

Oxidation number of O: \(-2\) → Total charge of all O’s: \((4 \times -2)\) \(-8\)

Oxidation number of P: \(+5\) → Total charge of all P’s: \(+5\)

Total charge of ion = \(-3\)
Identifying Redox Reactions

- In redox reactions, both oxidation and reduction occur together and one element is reduced while another is oxidized.
- In redox reactions the oxidation numbers of some atoms change.
Assign oxidation numbers to ALL elements in an equation and find the change

+4 -2   +1 -1   +2 -1   +1 -2   0
-2
-2       -1       +1

$\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + 2\text{H}_2\text{O} + \text{Cl}_2$
Spectator ions

- are ions whose oxidation number remains unchanged during the course of a redox reaction. They are generally not included in the half-reactions

\[
\begin{align*}
+4 & \quad -2 & \quad +1 & \quad -1 & \quad +2 & \quad -1 & \quad +1 & \quad -2 & \quad 0 \\
-2 & \quad +1 & \quad -1 & \quad +1 \\
\end{align*}
\]

\[\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + 2\text{H}_2\text{O} + \text{Cl}_2\]

Oxygen and Hydrogen are spectator ions
Writing Half Reactions

- think of a redox reaction as consisting of two half-reactions.
- One half-reaction represents the loss of electrons (oxidation)
- the other half-reaction represents the gain of electrons (reduction)
- Balance each half-reaction for both mass and charge
Example

$\text{MnO}_2 + 4\text{HCl} \rightarrow \text{MnCl}_2 + 2\text{H}_2\text{O} + \text{Cl}_2$

The half-reactions are:

Reduction: $\text{Mn}^{4+} + 2\text{e}^{-} \rightarrow \text{Mn}^{2+}$

Oxidation: $2\text{Cl}^{-} \rightarrow \text{Cl}_2 + 2\text{e}^{-}$
Identifying Agents in Redox Rxns

- **Oxidizing agent** is any element that in a redox reaction is reduced or gains electrons. It causes another element to lose electrons (to be oxidized).
- Strong oxidizing agents – \( \text{O}_2 \), halogens and \( \text{H}_2\text{O}_2 \)
- **Reducing agent** is any element that in a redox reaction is oxidized or loses electrons. It causes another element to gain electrons (to be reduced).
- Strong reducing agents Group 1 metals, hydrogen and carbon

*THE “AGENTS” WILL ALWAYS BE REACTANTS*
Using Table J…….

- Group 1 Metals are **strong reducing agents**, but **weak oxidizing agents**. (Look at their positions on Table J)

- Au, Ag & Cu are **strong oxidizing agents**, but **weak reducing agents**. (Look at their positions on Table J)

- F₂ is a **strong oxidizing agent**, but a **weak reducing agent**. (Look at its position on Table J)

- I₂ is a **weak oxidizing agent**, but a **strong reducing agent**. (Look at its position on Table J)
Reference **Table J** lists the activity of metals and non-metals.

An element on the list is more reactive than any element below it.

The **higher** the neutral **metal** elements (**atom**) will be oxidized by metal **ions** below them.

The **higher** neutral **non-metals** (**atom**) will be reduced by the non-metal **ions** below them.

In a **single replacement reaction**, a higher element (**atom**) will replace a lower element (**ion**) in a compound.
Examples

- Ex. #1 Will Ba react with Mn$^+$? Yes, b/c Mn$^+$ is below Ba

- Ex. #2 Will Na$^+$ react with Cr?  No, b/c Na$^+$ is above Cr

- Ex. #3 What are the products of the reaction Mg + Co(NO$_3$)$_2$ →? Mg is higher than Co. ∴ it can replace it
Consider the half reactions:

\[
\begin{align*}
\text{Al}^0 & \rightarrow \text{Al}^{3+} + 3e^{-1} \\
\text{Cl}_2^0 & + 2e^{-1} \rightarrow 2\text{Cl}^{-1}
\end{align*}
\]

- Before the half-reactions can be added, it is first necessary to conserve charge.
- That is, the number of electrons lost must equal the number of electrons gained, or the total charge must equal zero.
Multiplying the entire oxidation reaction by 2 and multiplying the entire reduction reaction by 3 and then adding both reactions gives the complete redox reaction.

\[
2(Al^0 \rightarrow Al^{3+} + 3e^{-1}) \\
3(Cl_2^0 + 2e^{-1} \rightarrow 2Cl^{1-}) \\
2Al + 3Cl_2 \rightarrow 2Al^{3+} + 6Cl^{1-} \ *
\]

*This is called a Net Ionic Equation- It has no spectator ions, no electrons, just the substances that are reduced and oxidized. It is balanced for mass and charge.*
Balancing Redox Rxns

1. Assign oxidation numbers to each element
2. Find Redox
3. Write the oxidation and the reduction half-reactions
4. Balance the electrons gained and the electrons lost by multiplying each half rxn with coefficients which can be used as coefficients to start to balance the equation
5. Inspect
Balance the following equation:

\[ \text{Cu} + \text{HNO}_3 \rightarrow \text{NO} + \text{Cu(NO}_3\text{)}_2 + \text{H}_2\text{O} \]
Balance the following equation:

Assign oxidation numbers:

\[
\begin{array}{cccccc}
0 & +1 & +5 & -2 & +2 & -2 & +2 & +5 & -2 & +1 & -2 \\
Cu & + & HNO_3 & \rightarrow & NO & + & Cu(NO_3)_2 & + & H_2O \\
\end{array}
\]
Find Redox and write half rxns

\[ \text{RED: } \text{Cu}^0 \rightarrow \text{Cu}^{+2} + 2e^- \]

\[ \text{OX: } \text{N}^{+5} + 3e^- \rightarrow \text{N}^{+2} \]
Balance electrons gained and electrons lost

\[
\begin{array}{cccccc}
0 & +1 & +5 & -2 & +2 & -2 & +2 & +5 & -2 & +1 & -2 \\
\end{array}
\]

\[
\text{Cu} + \text{HNO}_3 \rightarrow \text{NO} + \text{Cu(NO}_3\text{)}_2 + \text{H}_2\text{O}
\]

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

\[
2(\text{N}^{+5} + 3e^- \rightarrow \text{N}^{+2})
\]
Inspect

\[ 3\text{Cu} + 8\text{HNO}_3 \rightarrow 2\text{NO} + 3\text{Cu(NO}_3\text{)}_2 + 4\text{H}_2\text{O} \]
Electrochemistry

- the branch of chemistry that is the study of the relationship between electric forces and chemical rxns.
Chemistry Meets Electricity

- Electrochemical cells (Voltaic or Galvanic cells) - a device that can change chemical energy into electrical energy

- They are called batteries
Electrochemical rxns in a battery cause a greater electron density in the negative terminal than in the positive terminal.

Electrons repel each other, so there is a higher “pressure” on the electrons in the negative terminal, which drives the electrons out of the battery and through your cell phone.

Electrical “pressure” is often called electrode potential or voltage and is expressed in volts.
Components of Electrochemical (Voltaic, Galvanic) Cell

- consists of two metallic electrodes (an electrode is a conductor that connects with a nonmetallic part of a circuit) separated by an electrolyte (a solution that conducts electricity).
- The **anode** is where oxidation takes place ANOX; VAN (Voltaic (cell) Anode Negative)
- The **cathode** is Where Reduction Occurs RED CAT. The cathode is positive.
- **FAT CAT** (electrons) – Flow Anode to Cathode
- Pathways for moving charges – a **salt bridge** allows for the migration of ions
Electrochemical Cell

*Start with Table J

Mnemonic Device: ANOX REDCAT
FATCAT-electrons Flow Anode to Cathode

<table>
<thead>
<tr>
<th>Oxidation</th>
<th>Reduction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zn$^0$ → Zn$^{+2}$ + 2e$^-$</td>
<td>Cu$^{+2}$ + 2e$^-$ → Cu$^0$</td>
</tr>
</tbody>
</table>

Strong Reducing Agent: 0.76 V

Strong Oxidizing Agent: 0.34 V

Net Ionic Equation: Zn$^0$ + Cu$^{+2}$ → Zn$^{+2}$ + Cu$^0$

Over all cell voltage: +1.10 V = Spontaneous
Electrolytic cells requires electrical energy to produce chemical changes. This process is called electrolysis.

- **Signs of the electrodes switch!**
- **Still ANOX REDCAT**
- **Anode is positive ; cathode is negative**
Electroplating

- the object to be plated is made the **cathode**. The solution in which the object is submerged contains ions of the metal to be plated, and the anode is a piece of the same metallic element. At the cathode reduction causes the metallic ions to come out of solution onto the surface to be plated. As these ions are removed from solution at the cathode, they are replaced at the anode by oxidation of the same element.

- **Plate the Red Cat** – the cathode is where the object to be **plated** is and where **reduction** takes place
The penny is attached to the cathode and the gold bars are the anode.
you can plate a metal onto another as long as the plating metal is below the metal you want to plate on.

Example - Ag can be plated on any metal above, but not on Au because Au is below it!
Cell Comparison

**Electrolytic**
1. non-spontaneous-needs an outside source of energy the overall cell potential $E^0$ is negative
2. ANOX REDCAT
3. anode is +; cathode is -
4. Endothermic b/c of outside source of energy needed to operate the cell; energy is on the reactant side of equation
5. Electrical energy is need to create the chemical energy
Example; recharging a battery; cell phone charging

**Voltaic/Galvanic (Electrochemical)**
1. spontaneous- does not need an outside source of energy the overall cell potential $E^0$ is positive
2. ANOX REDCAT
3. anode is -; cathode is +
4. Exothermic because energy is created; energy will be on Product side of equation
5. Chemical energy is converted to electrical energy
Example; a battery