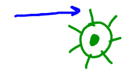
# CHAPTER 20

Oxidation-Reduction Reactions

## 20.1 The Meaning of Oxidation and Reduction

Oxidation reactions are the principal source of energy on earth.



## Redox Reactions

*Also known as…*

**Oxidation-Reduction Reactions: chemical changes that occur when electrons are transferred between reactants.**

- Oxidation originally meant the combination of an element with oxygen to give oxides.

  $2Na + O_2 \rightarrow 2Na_2O$

- Over the years, reduction has meant the loss of oxygen from a compound.

  $2Fe_2O_3 \rightarrow 4Fe + 3O_2$

## Oxidation Numbers

**An oxidation number is a positive or negative number assigned to an atom according to a set of arbitrary rules.**

**Tracks the movement of electrons**

## Rule #1

**Uncombined element = 0**

- **Diatomic = 0.** ***(Non polar)***

- $N_2$ Fe Au $O_2$ $C_4$

# Rule #2

- Monoatomic Ion
- standard Ionic charge
- There is a difference between Br & $Br^-$

- $Br^-$ $Fe^{3+}$ $Mg^{2+}$

# Neutral Compounds

NaCl

- In ionic compounds, use the charges as their oxidation numbers. *(look for metal)*

- In molecular compounds, the charges are the result of the electrons transferred.

$NH_3$ $H_2O$

# Rule #3

Nonmetals

Molecule = covalent

- In a molecule, the more electronegative element has the oxidation number that it would have if it were in an ionic compound.

- $NH_3$ $CO_2$ $H_2O$

# Rule #4

- Fluorine always has an oxidation number of -1 since it is the most electronegative element.

# Rule #5

Oxygen is usually –2.

*(Unless in peroxide, then it is -1: $H_2O_2$)* *

+2 −1

*(Unless bonded to F, then it is +2: $OF_2$)* *

$Na_2O$ $CO_2$ $P_2O_5$ $SO_2$

# Rule #6

Hydrogen in a compound

- +1 unless in a metal hydride like NaH, which is -1.
- Metal hydride is metal bonded *directly* to hydrogen.

- $H_2O$ NaOH KH $NH_4^+$ $MgH_2$
- Hydrogen bonded to a polyatomic ion is not a metal hydride!

# Rule #7

**Metals in groups 1A, 2A, & Al in group 3A form compounds in which the metal atom always has positive oxidation equal to their number of valence electrons.**

1-A= +1

2 A= +2

Al = +3

# Rule #8

**In neutral compounds, the sum of oxidation numbers = 0.**

$NaCl$   $Mg_3N_2$   $K_2S$   $S_2O_3$   $S_2O_4$

| | | |
|---|---|---|
| $CaO$ | $Ca_3(PO_4)_2$ | $H_2SO_4$ |
| $H_2$ | $O_2$ | $SCl_6$ |
| $BaCl_2$ | $NaH$ | $KOH$ |
| $CaSO_4$ | $LiNO_3$ | $CO_2$ |
| $H_2O_2$ | $CH_4$ | $Na_2CO_3$ |

# Rule #9

**Charged Compounds or Polyatomic ions, the sum of the oxidation # equals the charge of the polyatomic ion.** ***(Not Neutral!!!)***

$SO_4^{2-}$   $MnO_4^-$   $NH_4^+$   $NO_3^-$   $PO_4^{3-}$

## Oxidation Number Review

$PbO_2$   $HSO_4^-$   $MgF_2$   $Al(NO_3)_3$

$Co_2(SO_4)_3$

## Oxidation

**The combination of an element or compound with oxygen to give oxides.**

- $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
- $C + O_2 \rightarrow CO_2$
- $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

## Reduction

- The loss of an oxygen from a compound. The amount of solid material has decreased.
- Iron oxide is reduced to iron by losing an electron.
- Carbon is oxidized to $CO_2$ by gaining oxygen.

$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$

- Oxidation and reduction occur simultaneously.
- No oxidation occurs without reduction and no reduction occurs without oxidation.

## Electron Shift in Redox Reactions

- Definition has expanded:
- Oxidation is complete or partial loss of electrons or gain of oxygen
- Reduction is complete or partial gain of electrons or loss of oxygen

LEO the lion goes GER

(Lose electrons oxidize) (Gain electrons reduce)

- Reducing Agent loses electrons
- Oxidizing Agent accepts electrons

## "LEO the lion goes GER"

$Mg + S \rightarrow MgS\ (Mg^{2+} + S^{2-})$

**Oxidation:**

$Mg \rightarrow Mg^{2+} + 2e^-$ (loss of $e^-$)

**Reduction:**

$S + 2e^- \rightarrow S^{2-}$ (gain of $e^-$)

Reducing Agent: substance in a redox reaction that donates electrons. (Mg)

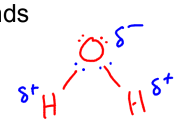
Oxidizing Agent: substance in a redox reaction that accepts electrons. (S)

$Mg + S \rightarrow MgS$

Oxidation and reduction also occurs in molecular substance that don't have a charge. These compounds use covalent bonds

∅ ∅ +1 -2

$2H_2 + O_2 \rightarrow 2H_2O$




**Hydrogen is oxidized by bonding with oxygen due to <u>partial loss</u> of electrons.**

## OXIDATION NUMBERS

**<u>Ionic compound</u>:**

**NaCl ➔ $Na^{+1}$ and $Cl^{-1}$**

**<u>Molecular Compound</u>:**

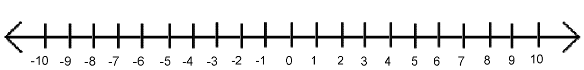
**$H_2O$ ➔ $H^{+1}$, $H^{+1}$, $O^{-2}$**

**Mg + S → MgS**

| Oxidation | Reduction |
|---|---|
| Loss of Electrons | Gain of Electron |
| Gain of Oxygen | Loss of Oxygen |
| Increase in Oxidation #. | Decrease in Oxidation #. |
| $Mg \rightarrow Mg^{2+} + 2e^-$ | $S + 2e^- \rightarrow S^{2-}$ |
| Reducing Agent = Mg | Oxidizing Agent = S |

## Oxidation Number Changes In Chemical Reactions

- **An increase in the oxidation number of an atom signifies oxidation.**
- **A decrease in the oxidation number of an atom signifies reduction.**
- **It is possible for the same atom to oxidize and reduce.**




$Ca + H_2SO_4 \longrightarrow CaSO_4 + H_2$

O: _____R: _____OA: _____RA: _____

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$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$

O: _____R: _____OA: _____RA: _____

**How can the same element oxidize and reduce? It is rare, but does it happen…**

$$2H_2O_2 \longrightarrow O_2 + 2H_2O$$

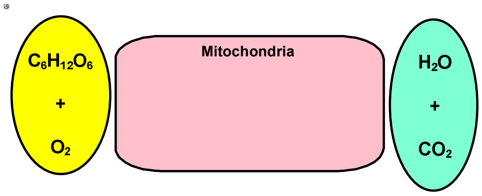
O: _____ R: _____ OA: _____ RA: _____

## 20.2 Balancing Redox Equations

**The total number of electrons gained in reduction must equal the total number of electrons lost in oxidation**

## Electron Transport Chain

http://vcell.ndsu.edu/animations/etc/movie-flash.htm




## Classes of Reactions

- **All other reactions in which electrons <u>are not</u> transferred.**
- **<u>Redox Reactions</u>: where electrons are transferred.**

## Determining Redox Reactions

- **Find oxidation numbers.**
- **See if they change?**
- **If they change, then it is a redox reaction.**

## Identifying Redox RXNS

- **Double replacement reactions and acid-base reactions are not redox reactions.**
- **Most other reactions are redox!**
- **Use oxidation numbers to determine if its a redox reaction.**

- **<u>Redox</u>: many single replacement, combination, decomposition, and combustion reactions.**
- **Double replacement and acid-base reactions are <u>not</u> redox reactions.**

## Which of these is a Redox reaction?

$CaSO_4 + Ba(NO_3)_2 \rightarrow Ca(NO_3)_2 + BaSO_4$

$2NaBr + Cl_2 \rightarrow 2NaCl + Br_2$

$H_2SO_4 + Mg(OH)_2 \rightarrow H_2O + MgSO_4$

- **Rules for oxidation number change method of balancing redox reactions**
- **MEMORIZE!!!**
- **Page 648**

## Oxidation Number Change Method for Balancing Redox Reactions

***Step 1*: Assign Ox Numbers**

***Step 2*: Determine what is oxidized and reduced**

***Step 3*: Use bracketed line to connect atoms undergoing oxidation and reduction and write the oxidation change on the line.**

- **Determine least common multiple between oxidation number change and reduction number change.**

***Step 4*: Place coefficients into the equation to reach the <u>number multiplied</u> to get the least common multiple.**

***Step 5*: Make sure the entire equation is balanced.**

$\_Mg + \_N_2 \rightarrow \_Mg_3N_2$

$\_Fe + \_O_2 \rightarrow \_FeO$

$\_PbS + \_O_2 \longrightarrow \_PbO + \_SO_2$

$\_C_2H_6 + \_O_2 \longrightarrow \_CO_2 + \_H_2O$

## Balancing Net Ionic Redox Reactions

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$$

$$Cu + 4H^+ + 4NO_3^- \rightarrow Cu^{2+} + 2NO_3^- + 2NO_2 + 2H_2O$$

## 20.3 Identifying Half-Reactions

Splitting a Redox reaction in half, focusing on the element that is oxidizing or reducing.

$$Fe + Cl_2 \rightarrow FeCl_3$$

Assign oxidation numbers and determine what oxidized and what reduced.

## Oxidation Half-Reaction

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$

***As products formed, Fe lost 3 electrons***

$Fe \rightarrow Fe^{3+} + 3e^-$ Half-Reaction

$2Fe \rightarrow 2Fe^{3+} + 6e^-$ Balanced Half-Reaction

## Reduction Half-Reaction

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$

***As products formed, $Cl_2$ gained 2 electrons***

$Cl_2 + 2e^- \rightarrow 3Cl^-$ Half-Reaction

$3Cl_2 + 6e^- \rightarrow 6Cl^-$ Balanced Half-Reaction

## Balancing Example
## Write the 2 Half-Reactions

$$Al + O_2 \rightarrow Al_2O_3$$

$$Al \rightarrow 2Al^{3+} + 6e^-$$

$$O_2 + 4e^- \rightarrow 3O^{2-}$$

## Balancing electrons in the Half-Reactions

$Al + O_2 \rightarrow Al_2O_3$

$4Al \rightarrow 4Al^{3+} + 12e^-$

$3O_2 + 12e^- \rightarrow 6O^{2-}$

## Balacing using Half-Reactions

1: Write net ionic equation and omit spectator ions

2: Write half-reactions from net ionic equation

3: Balance atoms and charges

4: Adjust coefficients so electrons lost and gained are equal

5: Add the half reactions and return spectator ions.

## Balacing using Half-Reactions - 1

$Fe + Cl_2 \rightarrow FeCl_3$

1. $Fe + Cl_2 \rightarrow Fe^{3+} + 3Cl^-$
2. $Cl_2 + 2e^- \rightarrow 3Cl^-$
   $Fe \rightarrow Fe^{3+} + 3e^-$
3. $3Cl_2 + 6e^- \rightarrow 6Cl^-$
   $Fe \rightarrow Fe^{3+} + 3e^-$
4. $3Cl_2 + 6e^- \rightarrow 6Cl^-$
   $2Fe \rightarrow 2Fe^{3+} + 6e^-$
5. $2Fe + 3Cl_2 \rightarrow 2FeCl_3$

## Balacing using Half-Reactions - 2

$Fe + S \rightarrow Fe_2S_3$

1. $Fe + S \rightarrow 2Fe^{3+} + 3S^{2-}$
2. $S + 2e^- \rightarrow 3S^{2-}$
   $Fe \rightarrow 2Fe^{3+} + 6e^-$
3. $3S + 6e^- \rightarrow 3S^{2-}$
   $2Fe \rightarrow 2Fe^{3+} + 6e^-$
4. $3S + 6e^- \rightarrow 3S^{2-}$
   $2Fe \rightarrow 2Fe^{3+} + 6e^-$
5. $2Fe + 3S \rightarrow Fe_2S_3$

## Balacing using Half-Reactions - 3

$K + Pb(SO_3)_2 \rightarrow Pb + K_2SO_3$

1. $K + Pb^{4+} + SO_3^{2-} \rightarrow Pb + 2K^+ + SO_3^{2-}$
2. $Pb^{4+} + 4e^- \rightarrow Pb$
   $K \rightarrow 2K^+ + 2e^-$
3. $Pb^{4+} + 4e^- \rightarrow Pb$
   $2K \rightarrow 2K^+ + 2e^-$
4. $Pb^{4+} + 4e^- \rightarrow Pb$
   $4K \rightarrow 4K^+ + 4e^-$
5. $4K + Pb(SO_3)_2 \rightarrow Pb + 2K_2SO_3$

## Chapter 20 Exam

40 Multiple Choice

3 Balancing with Oxidation Change Method

2 Balancing with Half-Reaction Method