SCH 4U

PART 1: OXIDATION-REDUCTION REACTIONS

- OXIDATION REACTIONS – Formally known as reactions involving oxygen.
- REDUCTION REACTIONS – Formally known as a chemical process of extracting metals from their ores, thus reducing the mass of the material from which they came.

RECENT DEFINITIONS:

- OXIDATION: loss of electrons from one substance to another
- REDUCTION: substance that gains the electrons lost by another substance
- Oxidation and reduction must happen together – also known as ***REDOX*** reactions.

Consider the reaction of magnesium metal with chlorine gas:

$$2\,Mg_{(s)} + Cl_{2\,(g)} \rightarrow 2\,MgCl_{2\,(s)}$$

$Mg \longrightarrow Mg^{2+} + 2e^-$ [Loss of Electrons is Oxidation]

$Cl_2 + 2e^- \longrightarrow 2Cl^-$ [Gain in Electrons is Reduction]

- The substance that is oxidized is the **reducing agent**. A reducing agent causes reduction to occur in the other substance
- The substance that is reduced is the **oxidizing agent**.

In the reaction of magnesium and chlorine, since the magnesium loses electrons, it is being oxidized and is called the reducing agent as it causes the chlorine atom to be reduced. Each chlorine atom gains an electron, so it is being reduced and is called the oxidizing agent as it causes the sodium atom to be oxidized.

Mg	+	Cl_2	$\longrightarrow$	$MgCl_2$
Loses e^- -- ***oxidized***		Gains e^- -- ***reduced***		
[reducing agent]		[oxidizing agent]		

- A redox reaction is a chemical reaction in which changes in oxidation numbers (or oxidation states) occur.

Many REDOX reactions are already well-known to us.

- Organic reactions
- Synthesis and decomposition reactions
- Many double displacement reactions

NOT ALL reactions are REDOX reactions – increase/decrease of oxidation states do not always occur.

EXAMPLE: Identify the oxidizing and reducing agents in the reaction of iron and bromine, producing iron (III) bromide.

SPONTANEITY of a REDOX REACTION:

- Will a reaction occur? – associated with single displacement reactions

Recall the Metal Activity Series from Gr. 11 Chemistry:
The metal activity series arranges metals in a column from most to least reactive, so that if a metal in the chart is located above the cation of a compound, the reaction will occur.

In tables of the Metal Activity Series, metals are often arranged from least reactive (weakest reducing agent) to most reactive (strongest reducing agent). The stronger the reducing agent, the greater the metal's ability to lose electrons to the other substance and form a new compound, or in other words, the greater the metal's ability to cause reduction (gain electrons) in the other substance, a reaction will occur. So when predicting if a REDOX reaction will occur, if the pure metal is located below the cation of the compound, a reaction will be spontaneous.

EXAMPLE: Which of the following reactions will proceed spontaneously?
If a reaction occurs, identify the oxidizing and reducing agents.

1. solid iron and aqueous tin (II) chloride

2. aqueous aluminum nitrate and copper metal

PART 2: OXIDATION NUMBERS

RULES FOR ASSIGNING OXIDATION NUMBERS (O.N.):

1. **Oxidation number of a free element is zero:**

2. **Oxidation number of monoatomic ion is equal to the charge on the ion:**

In compounds/polyatomic ions AND in the following order:

3. **fluorine has oxidation number -1.**

4. **hydrogen has oxidation number of +1, with the exception of metal hydrides where it has O.N. of -1**

5. **oxygen has oxidation number of -2.**

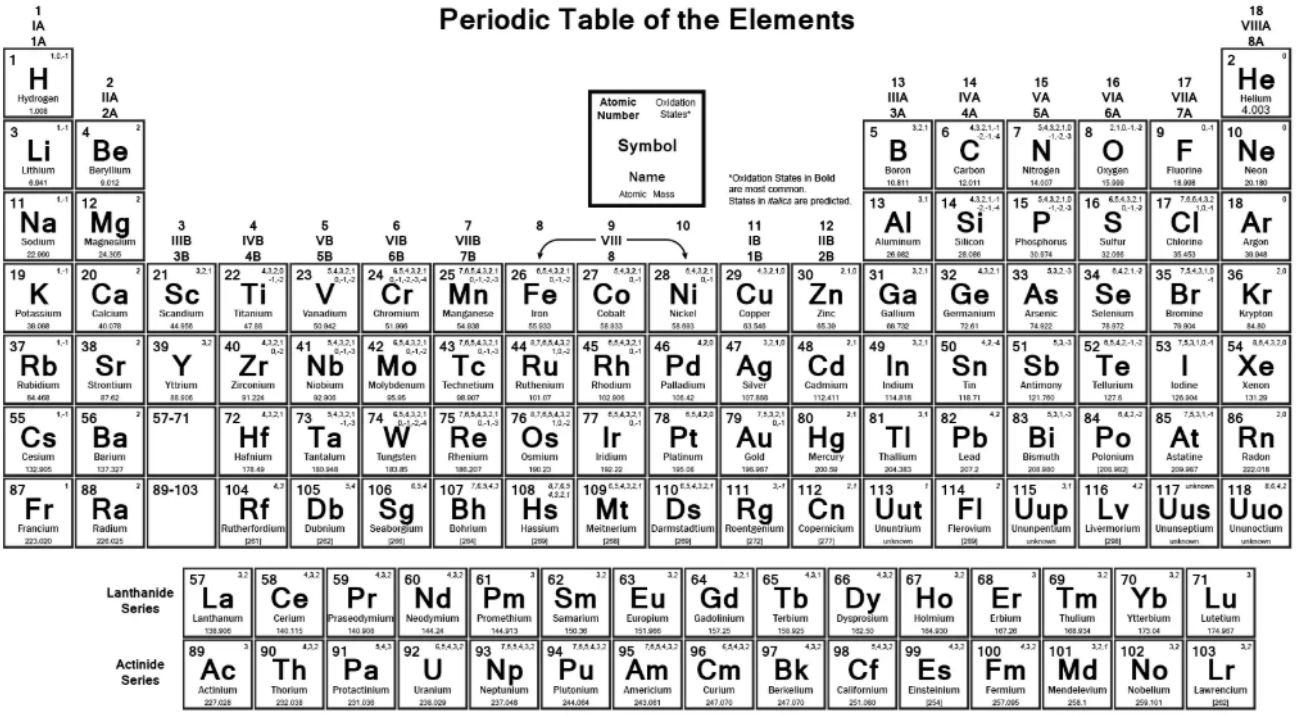
6. **molecules not containing H or O → the more electronegative element is assigned an O.N. equal to the negative charge it usually has when in an ionic compound.**

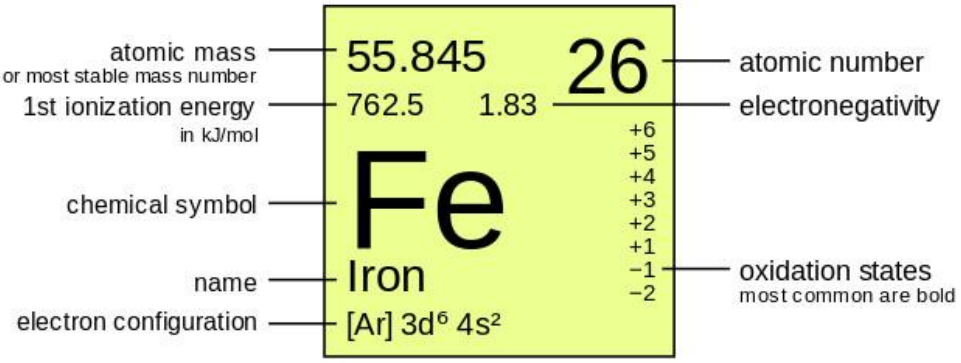
In general,

7. **Sum of oxidation numbers in a molecule is 0.**
   Sum of oxidation numbers in a polyatomic ion equals charge on particle.

EXAMPLES: Determine the oxidation number of each element in the given substance.

1. Cl_2O
2. Cl_2O_7
3. $HOCl$
4. $Mg_3(PO_4)_2$
5. NCl_3
6. HSO_4^-





Common Oxidation States of Nonmetals

Group 14	carbon	−4, +2, +4
Group 15	nitrogen	−3, +1, +2, +3, +4, +5
	phosphorus	−3, +3, +5
Group 16	sulfur	−2, +4, +6
Group 17	chlorine	−1, +1, +3, +5, +7
	bromine	−1, +1, +3, +5, +7
	iodine	−1, +1, +3, +5, +7

In addition to the values shown, atoms of each element in its pure state are assigned an oxidation number of zero.