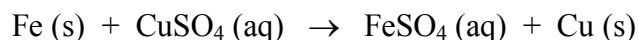
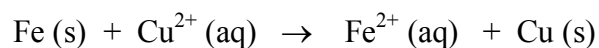


OXIDATION-REDUCTION REACTIONS

- **Oxidation-reduction** (redox) reactions are reactions that involve transfer of electrons from one species to another.
- As an example:



Net ionic equation:



- In this reaction, the electrons are transferred from iron atoms to copper ions to form iron ions and copper atoms.
- To better understand oxidation-reduction reactions, the concept of **oxidation numbers** was developed as a simple way of keeping track of electrons in a reaction.
- Using oxidation numbers, one can determine if electrons have been transferred between substances during a reaction.
- If transfer of electrons has occurred during the reaction, an oxidation-reduction reaction has occurred.

OXIDATION NUMBERS

- The oxidation number of an atom is the number of electrons lost, gained or unequally shared by an atom. Oxidation numbers can be zero, positive or negative.
- Oxidation number of a monatomic ion is equal to the charge of the ion. For example:



Na^+ = sodium has lost an electron O.N. = +1

Cl^- = chloride has gained an electron O.N. = -1

- In covalent compounds, the oxidation numbers are based on relative electronegativities.
- For non-polar covalent compounds each atom is assigned an oxidation number of 0 because they have the same electronegativity. For example:

	H—H		Cl—Cl	
EN	2.1	2.1	3.0	3.0
O.N.	0	0	0	0

- For polar covalent compounds, the element with the greater electronegativity is assigned the negative oxidation number, and the element with the lower electronegativity is assigned the positive oxidation number. For example:

	H—Cl	
EN	2.1	3.0
O.N.	+1	-1

- Many elements have multiple oxidation numbers. For example:

	N_2	N_2O	NO	N_2O_3	NO_2	N_2O_5	NO_3^-
O.N.. of N	0	+1	+2	+3	+4	+5	+5

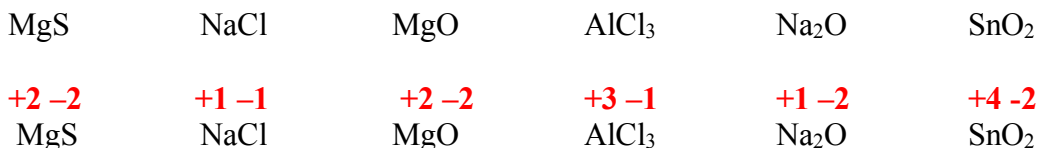
- The oxidation numbers for elements must be assigned based on a set of rules outlined next.

RULES FOR ASSIGNING OXIDATION NUMBERS
--

1. All elements in their free states have oxidation number of zero.



2. For monatomic ions, O.N. equals the charge of the ion.



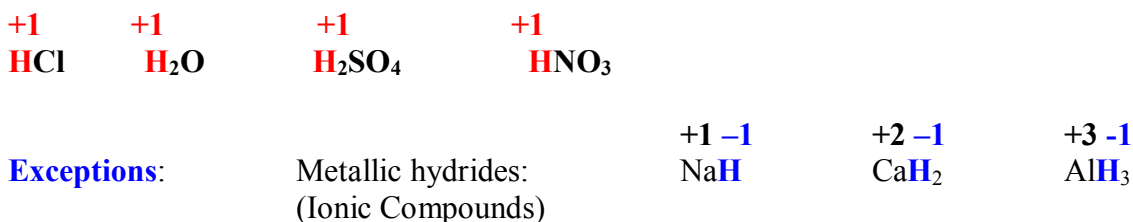
3. Oxygen is -2, except in peroxides, where it is -1 and OF₂, where it is +2.



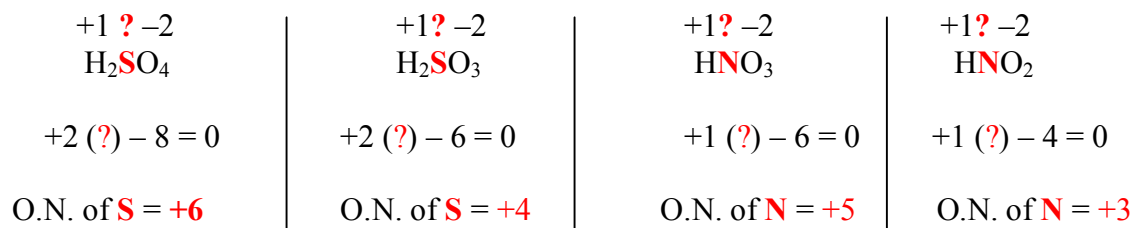
4. In Molecular Compounds, the negative ON is assigned to the atom that is more to the right and higher up on the Periodic Table



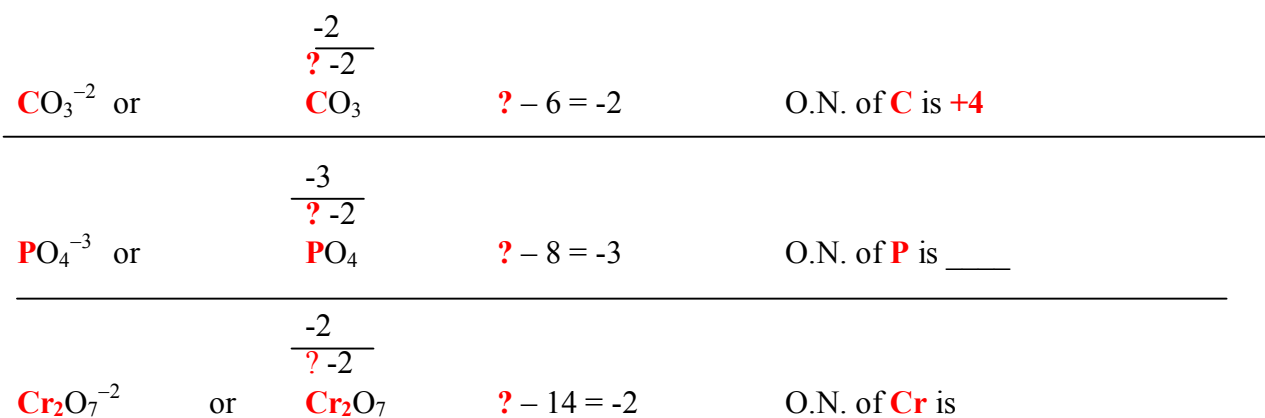
5. Hydrogen is +1, except when combined with metals, where it is -1.



6. The algebraic sum of the oxidation number of elements in a compound is equal to zero.

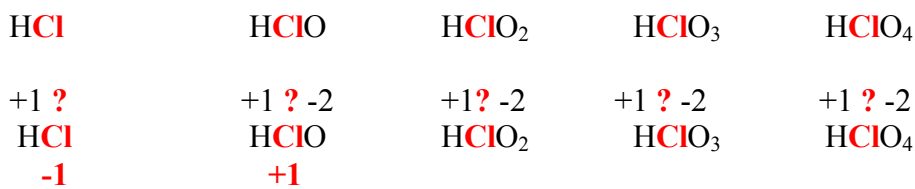


7. The algebraic sum of the oxidation numbers in a polyatomic ion is equal to the charge of the ion.



Examples:

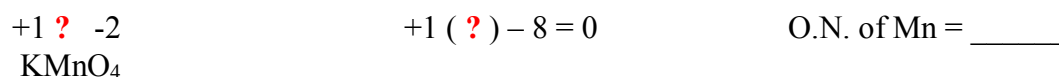
1. Calculate O.N. of **Cl** in the following Compounds:



OXIDATION NUMBER INCREASES

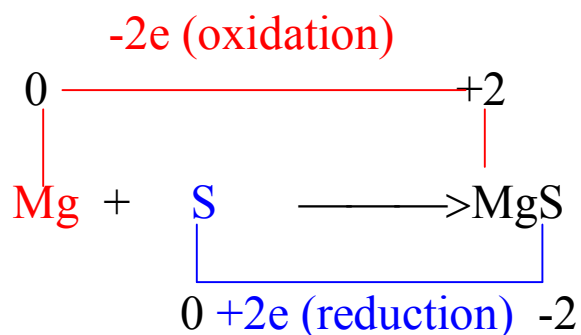


2. Calculate O.N. of Mn in KMnO_4



OXIDATION-REDUCTION REACTIONS
(Also referred to as Redox Reactions)

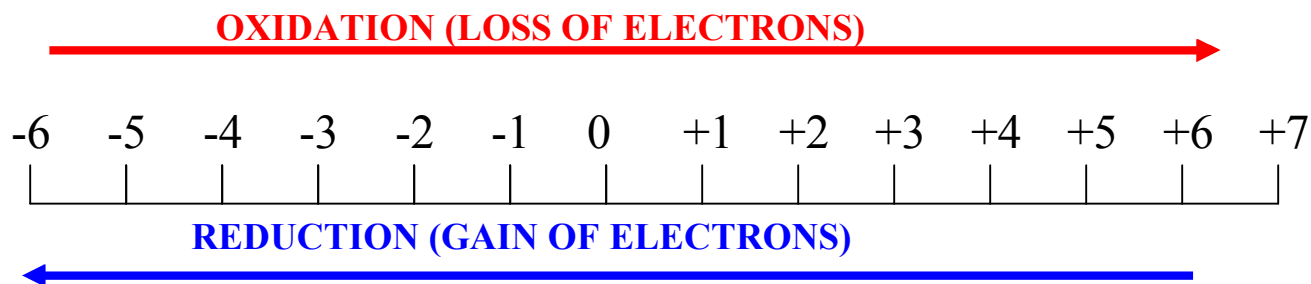
- Redox reactions are rxns in which the O.N.'s of at least 2 elements change

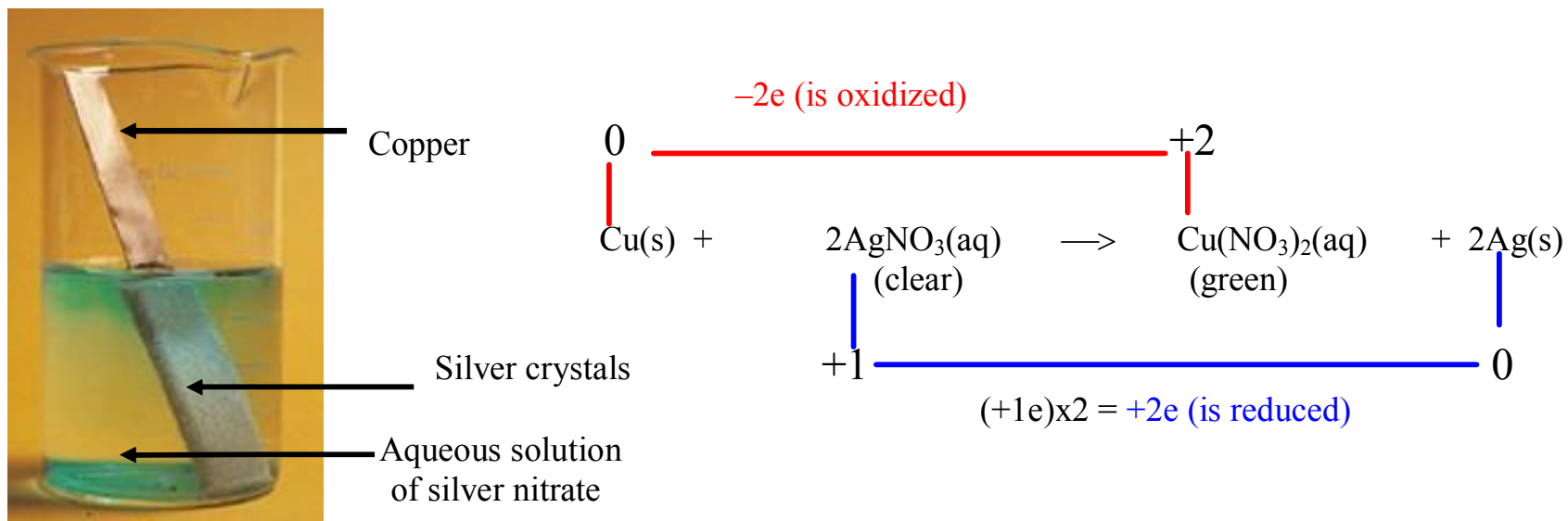
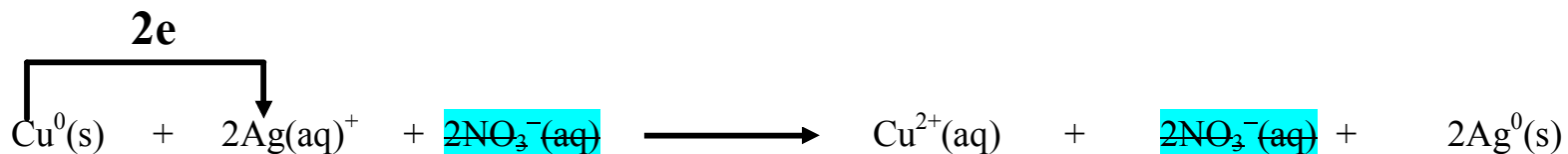
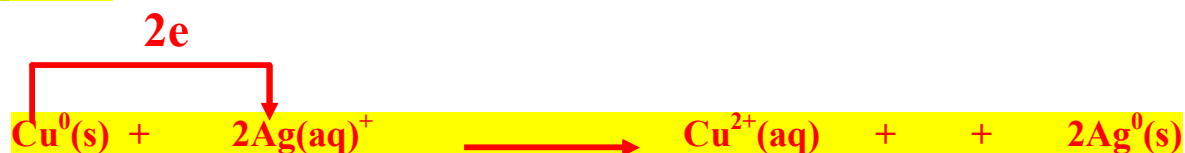

In General
An element that is oxidized: (Mg)

- loses electrons
- its O.N. increases
- is the Reducing Agent

An element that is reduced: (S)

- gains electrons
- its O.N. is reduced
- is the Oxidizing Agent



Example:**Total Ionic Equation:****Net Ionic Equation:**

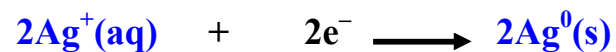
- Redox Reactions are discussed (sometimes balanced) by writing two Half – Reactions:

OXIDATION HALF-REACTION

- involves loss of electrons
- increase in Oxidation Number

**REDUCTION HALF-REACTION**

- involves gain of electrons
- decrease in Oxidation Number

**NOTE:**

Number of electrons lost in the
Oxidation Reaction

=

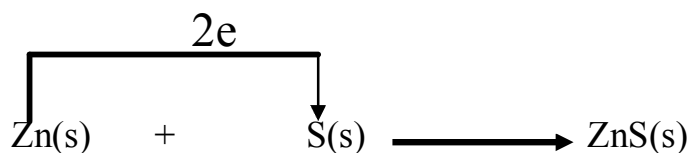
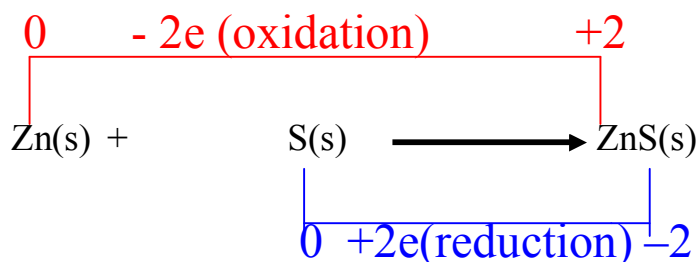
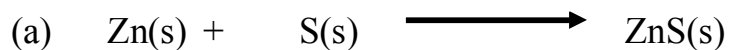
Number of electrons gained in the
Reduction Reaction

CONCLUSION:

Oxidation - Reduction (Redox) Reactions

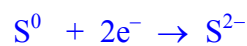
- are reactions in which the Oxidation Numbers of at least two elements change
- involve transfer of electrons:
 - from : the element that is oxidized (called Reducing Agent)
 - to: the element that is reduced (called Oxidizing Agent)

CLASSIFICATION OF SOME COMMON REDOX REACTIONS
1. Combination Reactions

Examples:

Oxidation Half-Reaction

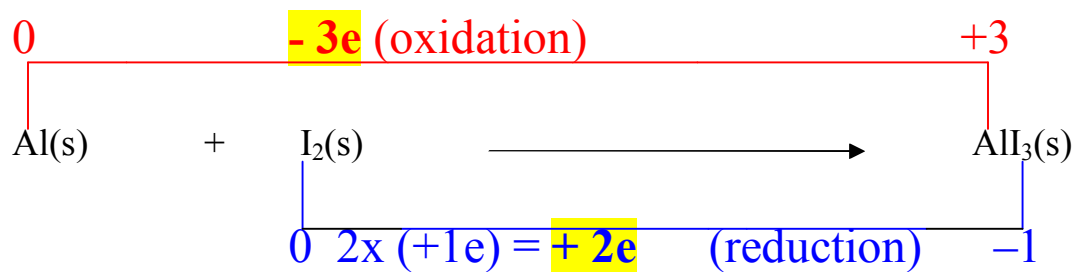
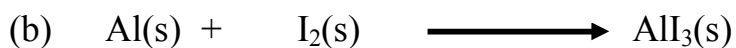

Zn is oxidized

Zn is the Reducing Agent

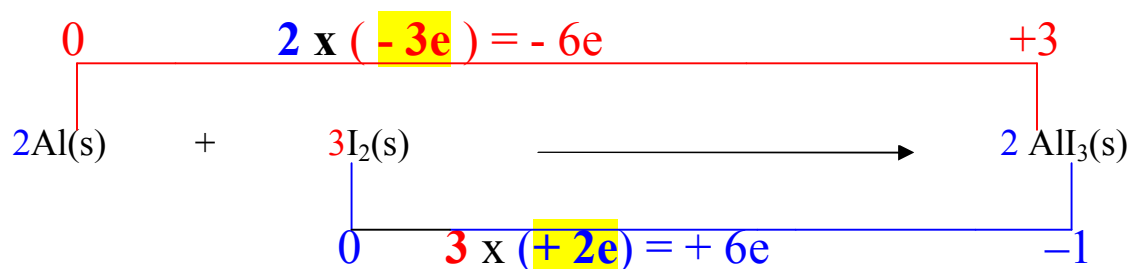
Reduction Half-Reaction


S is reduced

S is the Oxidizing Agent



NOTE: Number of electrons lost and gained must be equal:



Balancing can also be done by using the half-Reactions:

Oxidation Half-Reaction



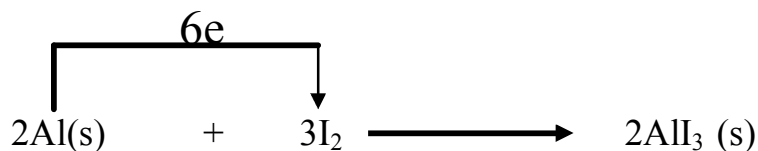
Reduction Half-Reaction



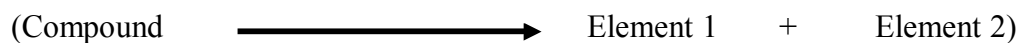
NOTE: Number of electrons lost and gained must be equal:



Balanced Redox Reaction:

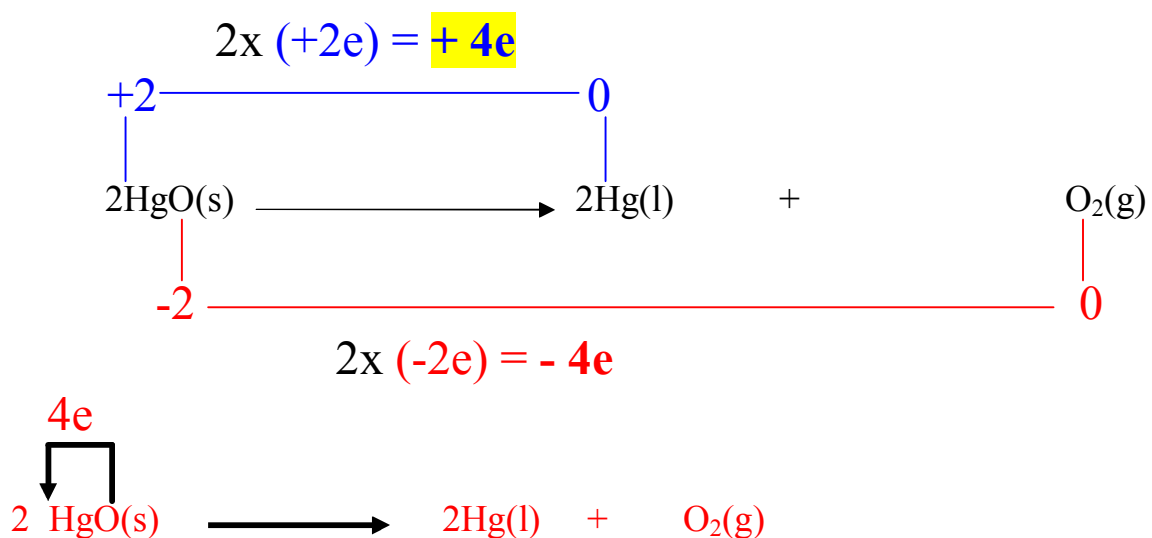
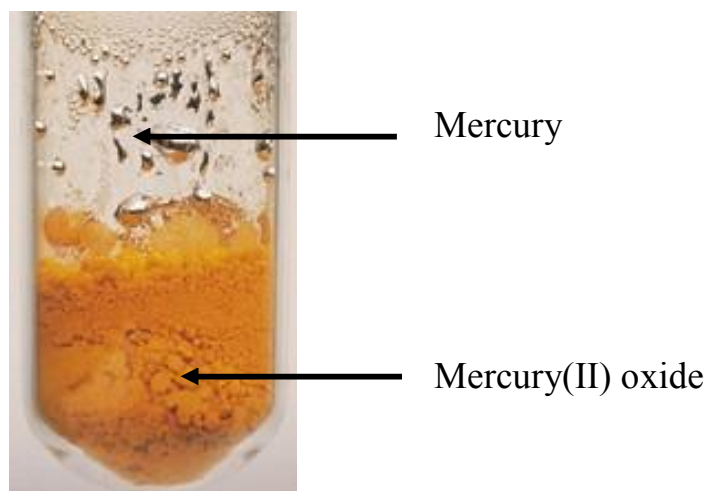


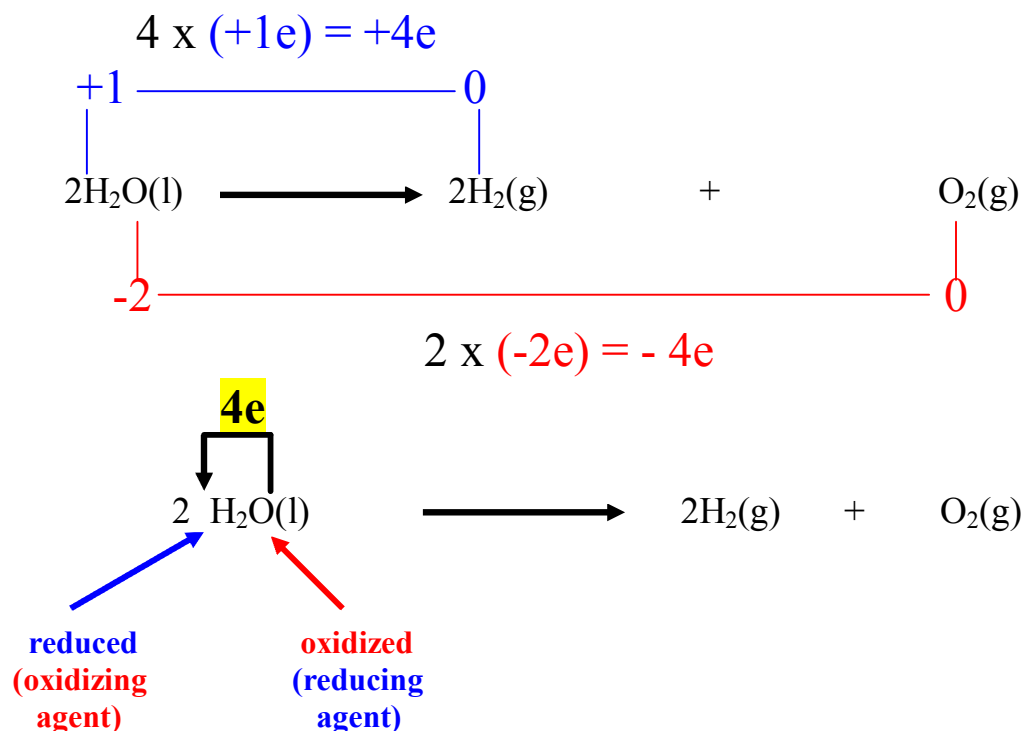
2. Decomposition Reactions



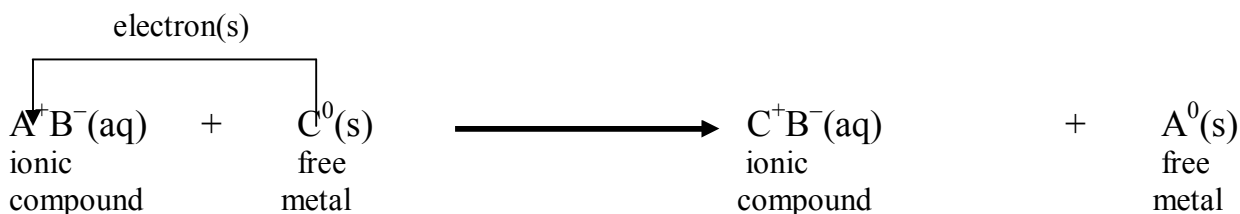
Examples:

(A) Red mercury(II) oxide decomposes when heated and produces silvery droplets of liquid mercury and oxygen gas.





3. Single Replacement Reactions



Oxidation Half-Reaction



C is oxidized

C is the Reducing Agent

Reduction Half-Reaction



A⁺ is reduced

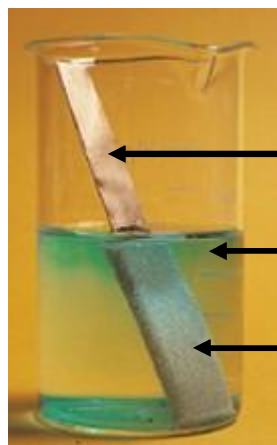
A⁺ is the Oxidizing Agent

Since metal C replaces A⁺ from its compound:

- C is more active than A
- C loses electrons easier than A
- C is a stronger reducing agent than A

Example 1

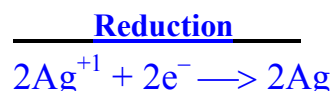
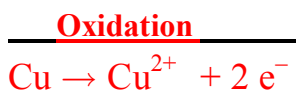
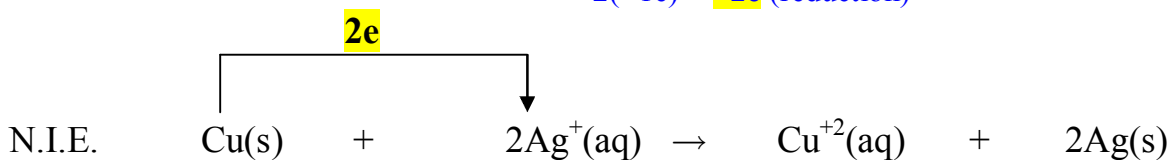
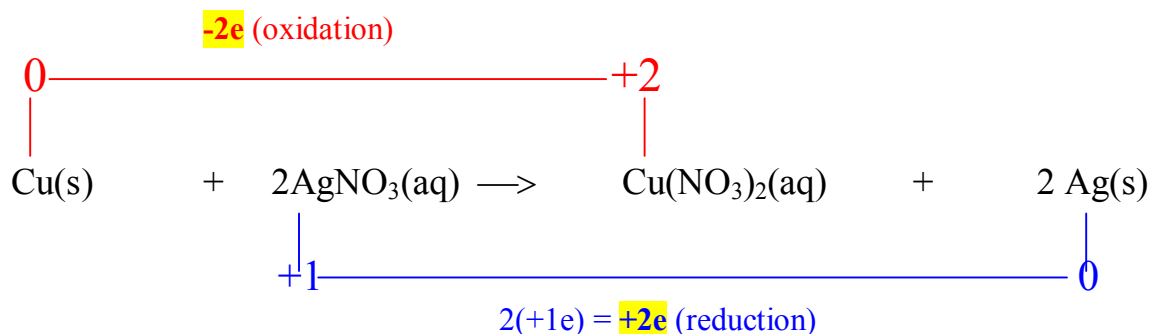
A copper strip dipped into a silver nitrate solution becomes coated with crystals of silver, while the solution turns greenish-blue, due to the presence of the copper(II) ion.



← copper strip

← copper(II) nitrate solution

← silver crystals



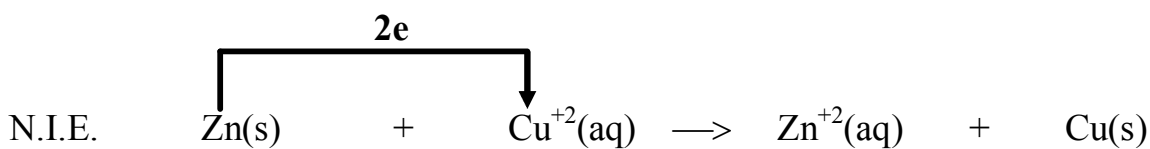
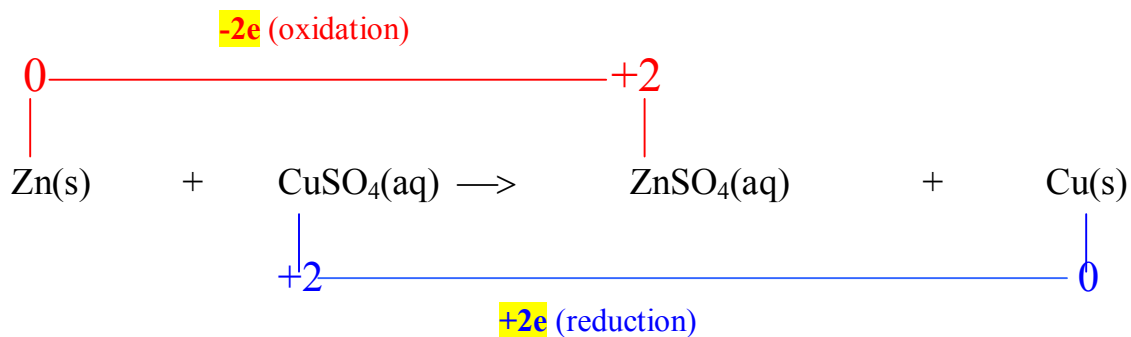
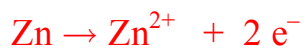
Since Cu replaces Ag^+ from its compound:

- Cu is more active than Ag
- Cu loses electrons easier than Ag
- Cu is a stronger reducing agent than Ag

It follows: $\text{Ag(s)} + \text{Cu(NO}_3)_2(\text{aq}) \longrightarrow \text{No Reaction}$

Since Ag does not replace Cu^{2+} from its compound:

- Ag is less active than Cu
- Ag does not lose electrons as easily as Cu
- Ag is a weaker reducing agent than Cu

Example 2:**Oxidation**

Zn is oxidized

Zn is the Reducing Agent

ReductionCu²⁺ is reducedCu²⁺ is the Oxidizing Agent

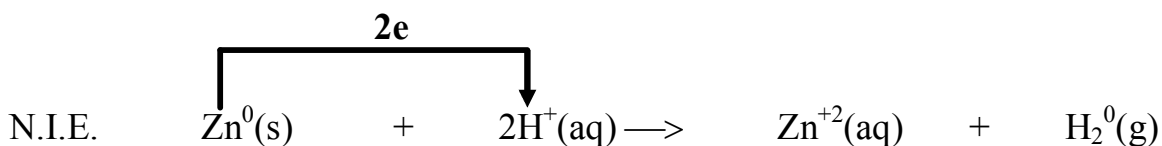
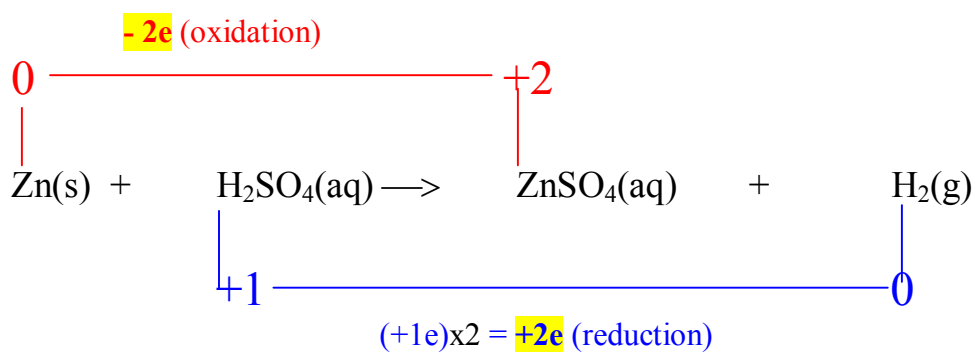
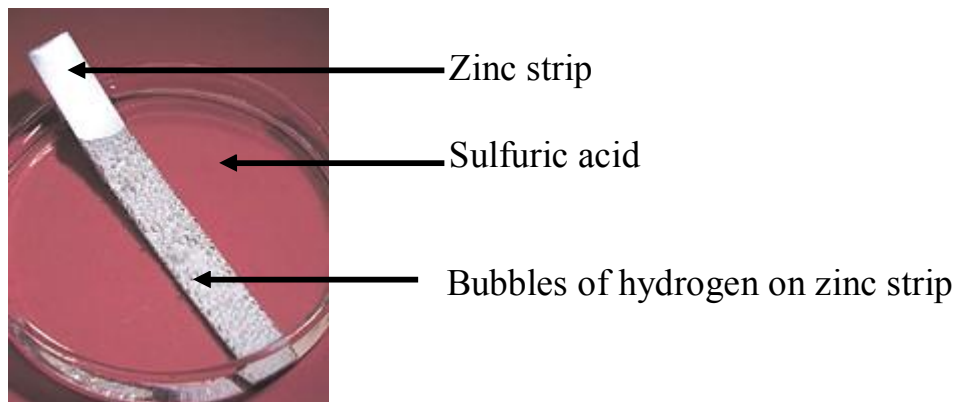
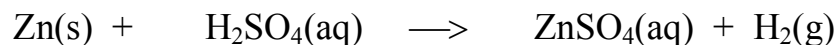
Since Zn replaces Cu²⁺ from its compound:

- Zn is more active than Cu
- Zn loses electrons easier than Cu
- Zn is a stronger reducing agent than Cu

It follows: $\text{Cu(s)} + \text{ZnSO}_4(\text{aq}) \longrightarrow \text{No Reaction}$

Since Cu does not replace Zn²⁺ from its compound:

- Cu is less active than Zn
- Cu does not lose electrons as easily as Zn
- Cu is a weaker reducing agent than Zn

Example 3:

Oxidation
 $\text{Zn} - 2\text{e} \longrightarrow \text{Zn}^{2+}$
 Zn is oxidized
 Zn is the Reducing Agent

Reduction
 $2\text{H}^+ + 2\text{e} \longrightarrow \text{H}_2^0(\text{g})$
 H^+ is reduced
 H^+ is the Oxidizing Agent

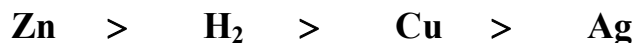
Since Zn replaces H^+ from its compound:

- Zn is more active than H_2
- Zn loses electrons easier than H_2
- Zn is a stronger reducing agent than H_2

Summing up:


The elements studied above may be ranked according to their activity:

Example 1:	Cu is more active than Ag	Cu > Ag
Example 2:	Zn is more active than Cu	Zn > Cu
Example 3:	Zn is more active than H ₂	Zn > H ₂



← ease of oxidation increases
metallic activity increases

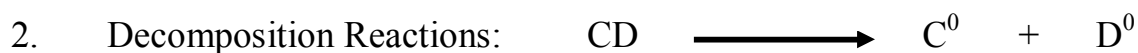
- Such a listing is called “An Activity Series”
- A more complete “Activity Series” includes the majority of common metals:

Ease of Oxidation increases 	K Ba Ca Na	<u>Very Active</u> : React with water to produce H ₂ Ex: $\text{Na(s)} + \text{H}_2\text{O(l)} \longrightarrow \text{NaOH(aq)} + \text{H}_2\text{(g)}$
	Mg Al Zn Cr Fe Ni Sn Pb	<u>Moderately Active</u> : React with nonoxidizing acids to produce H ₂ gas Ex: $\text{Zn(s)} + \text{H}_2\text{SO}_4\text{(aq)} \longrightarrow \text{ZnSO}_4\text{(aq)} + \text{H}_2\text{(g)}$
	H ₂	
	Cu Ag Hg Au	<u>Relatively inactive</u> : Do not react with nonoxidizing acids to produce hydrogen gas.

NOTE: A free metal can displace the ion of a second metal from solution if the free metal is above the second metal in the activity series.

SUMMARY OF OXIDATION REDUCTION REACTIONS

- Redox reactions can be identified by checking the Oxidation Numbers (O.N.) of all the elements involved.
- If the **O.N.'s of at least 2 elements change**, the reaction is an Oxidation – Reduction (Redox) Reaction.
- In redox reactions an **exchange of electrons** occurs between the reactants
- Redox reactions can be classified as:



3. Single Replacement Reactions:



4. Combustion Reactions: Burning in the presence of oxygen gas

SUMMARY OF TYPES OF CHEMICAL REACTION
I. DOUBLE DISPLACEMENT (METATHESIS) RXNS.

- The **ions** of the reactants are **exchanged**
- **Oxidation Numbers of the elements do not change**
(Charges of ions do not change from Reactant to Product side)



Double-Displacement Reactions can be further classified into:

1. Precipitation Reactions
2. Acid – Base Reactions
3. Reactions that form an unstable product

II. REDOX REACTIONS (Oxidation – Reduction Reactions)

- An **exchange of electrons** occurs between the reactants
- **Oxidation Numbers of at least 2 elements change**

Redox Reactions can be further classified into:

1. Combination Reactions
2. Decomposition Reactions
3. Single Replacement Reactions
4. Combustion Reactions