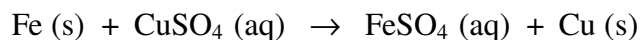
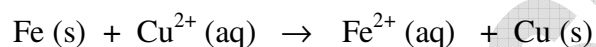


OXIDATION-REDUCTION REACTIONS

- **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 |

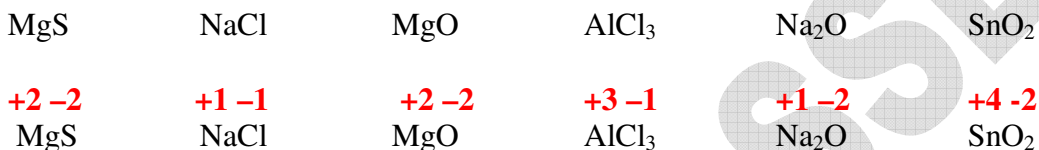
- The oxidation numbers for elements must be assigned based on a set of rules outlined next.

| |
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| RULES FOR ASSIGNING OXIDATION NUMBERS |
|--|

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



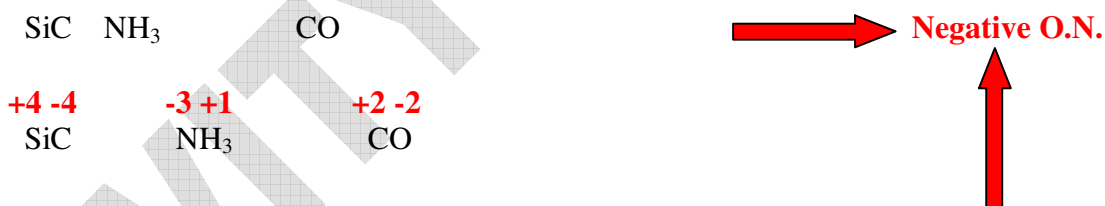
2. For monatomic ions, O.N. equals the charge of the ion.



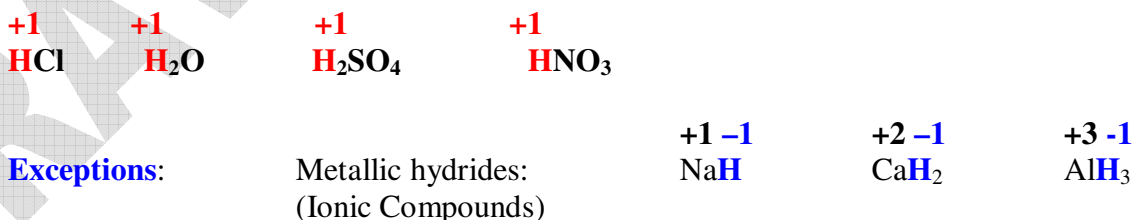
3. Oxygen is -2, except in peroxides, where it is -1 and OF₂, where it is +2.



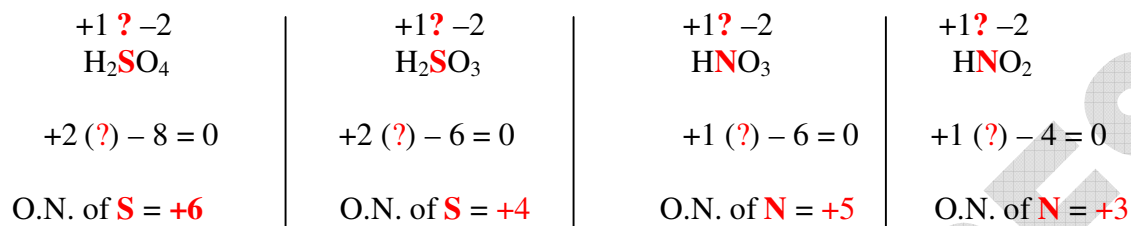
4. In Molecular Compounds, the negative ON is assigned to the atom that is more to the right and higher up on the Periodic Table



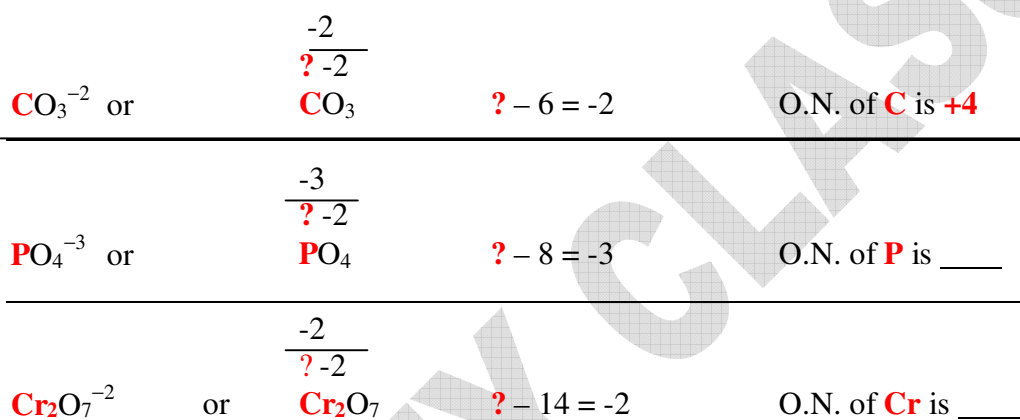
5. Hydrogen is +1, except when combined with metals, where it is -1.



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

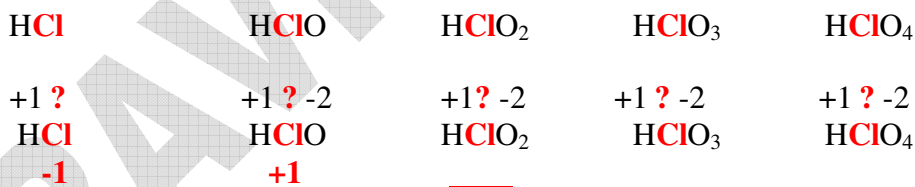


7. The algebraic sum of the oxidation numbers in a polyatomic ion is equal to the charge of the ion.



Examples:

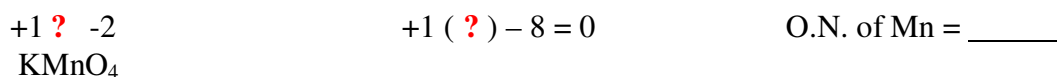
1. Calculate O.N. of **Cl** in the following Compounds:



OXIDATION NUMBER INCREASES

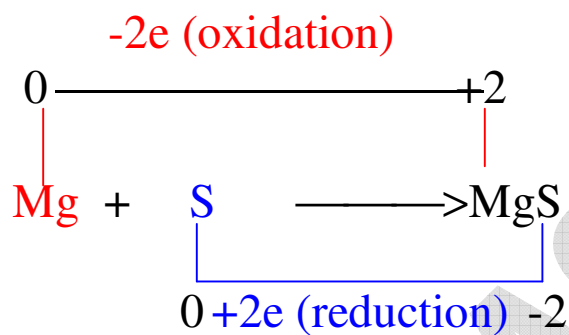


2. Calculate O.N. of Mn in KMnO_4



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

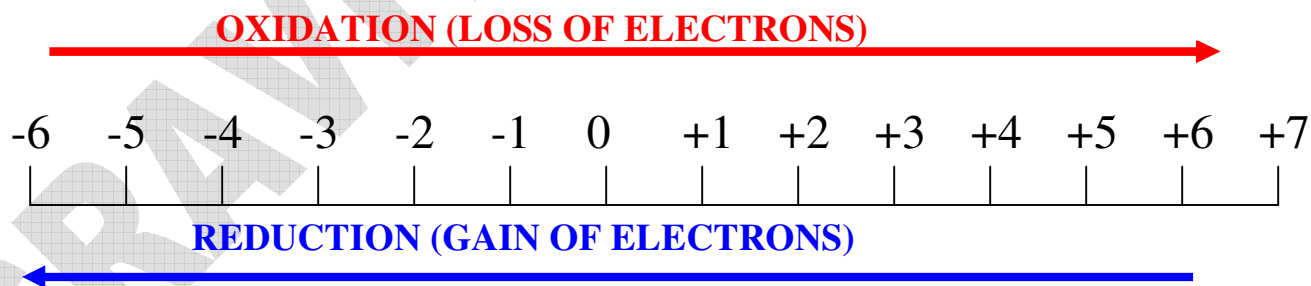
- Redox reactions are rxns in which the O.N.'s of at least 2 elements change

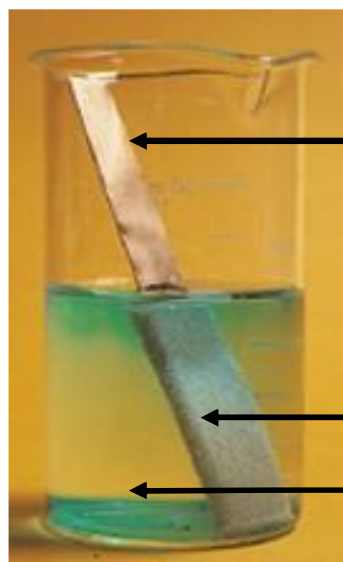

In General
Anelementthat is oxidized: (Mg)

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

Anelementthat is reduced: (S)

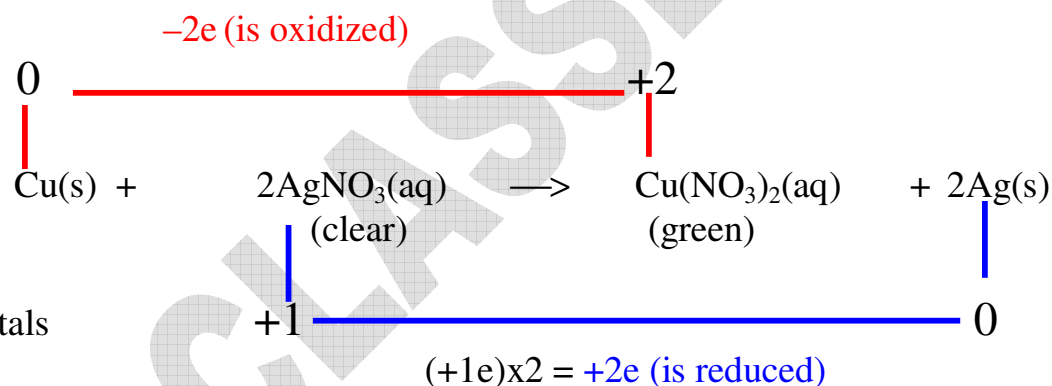
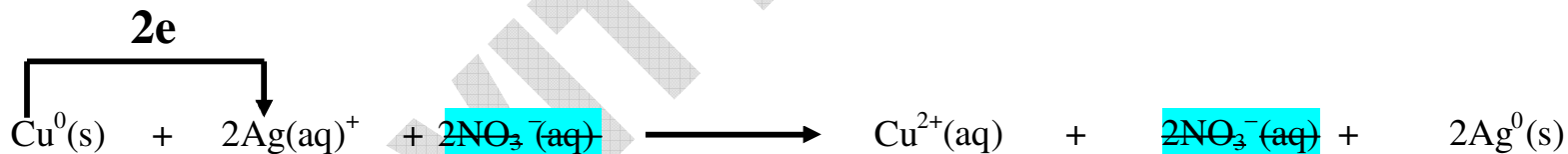
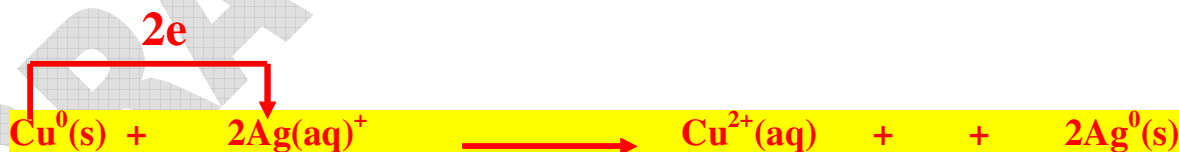
- gains electrons
- its O.N. is reduced
- is the Oxidizing Agent



Example:

Copper

Silver crystals

Aqueous solution
of silver nitrate**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

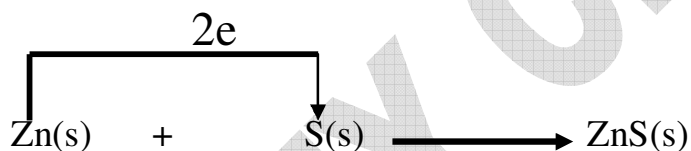
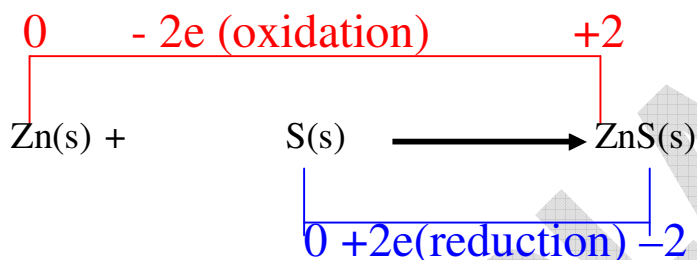
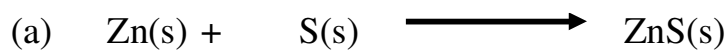
CONCLUSION:

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)

| |
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| CLASSIFICATION OF SOME COMMON REDOX REACTIONS |
|--|

1. Combination Reactions

Examples:

Oxidation Half-Reaction

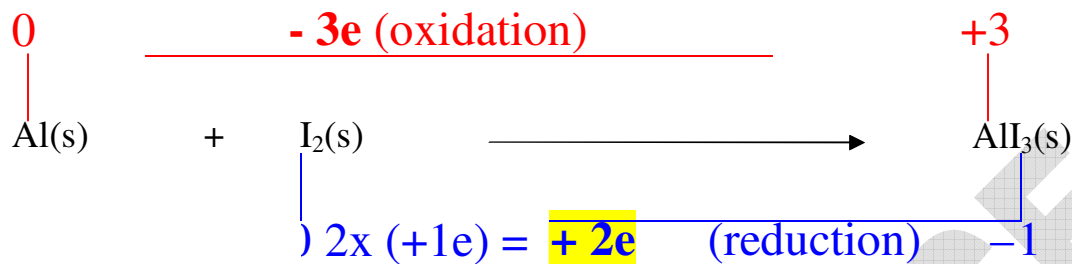
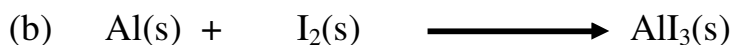

Zn is oxidized

Zn is the Reducing Agent

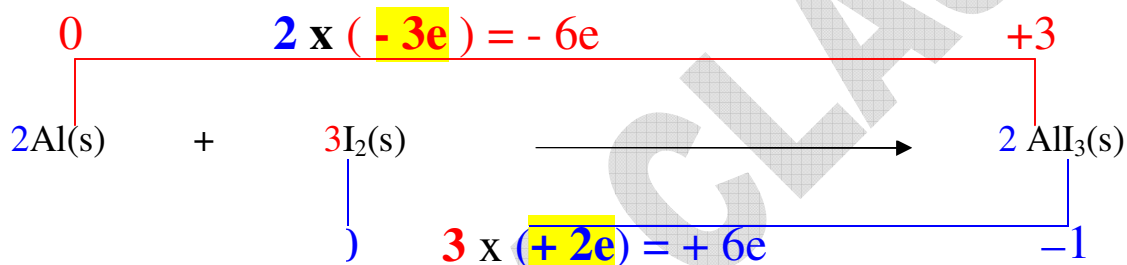
Reduction Half-Reaction


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

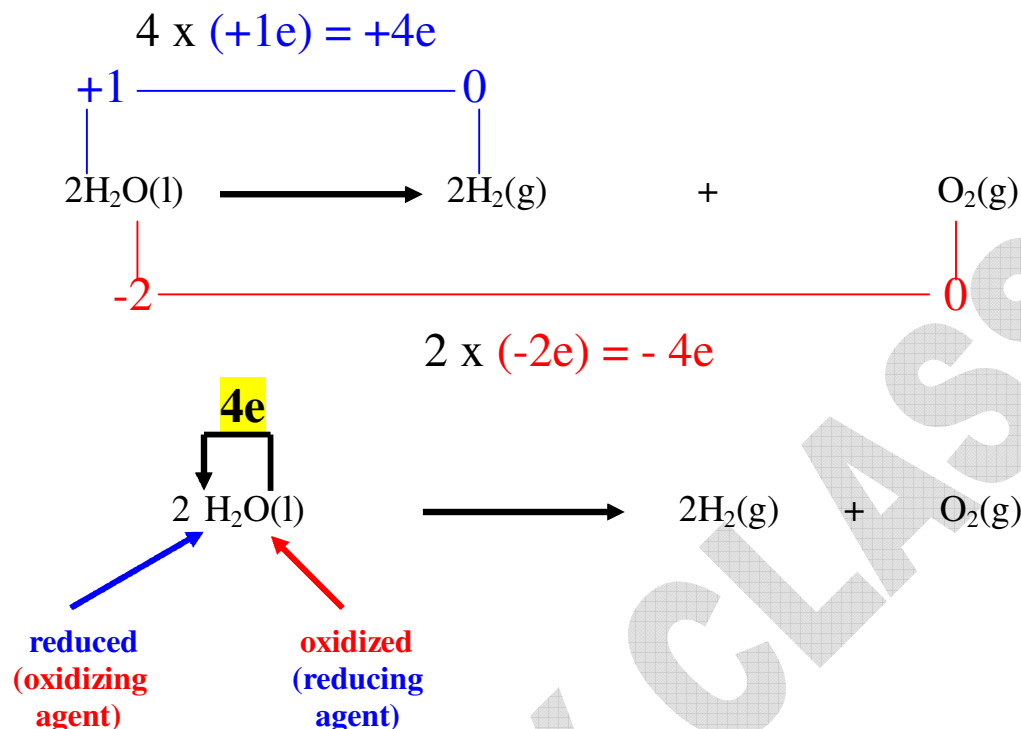
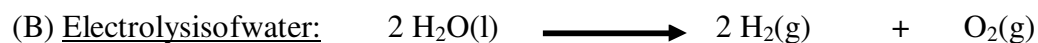


NOTE: Number of electrons lost and gained must be equal:

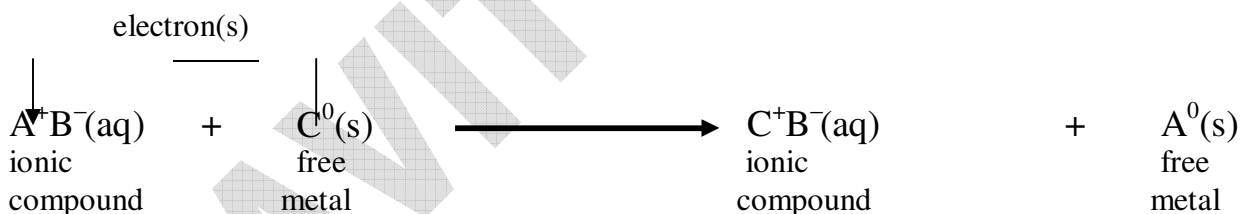


Balanced Redox Reaction:





3. Single Replacement Reactions



Oxidation Half-Reaction



C is oxidized

C is the Reducing Agent

Reduction Half-Reaction

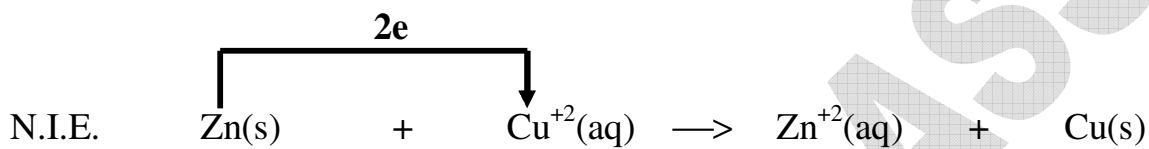
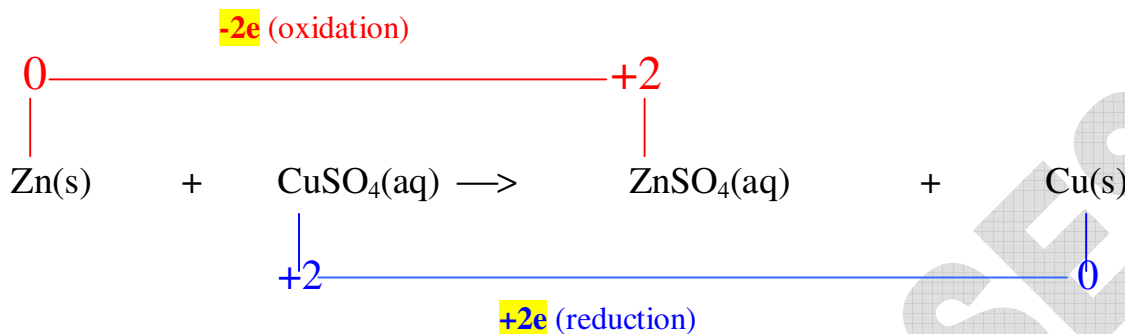
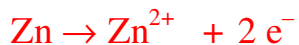


A⁺ is reduced

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 2:**Oxidation**

Zn is oxidized

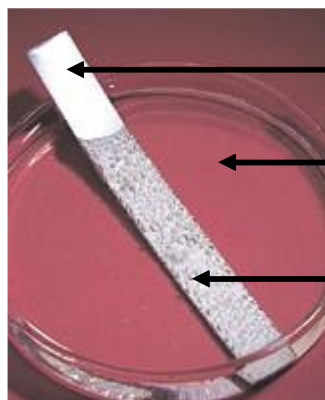
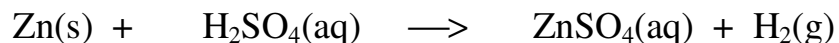
Zn is the Reducing Agent

ReductionCu²⁺ is reducedCu²⁺ is the Oxidizing AgentSince Zn replaces Cu²⁺ from its compound:

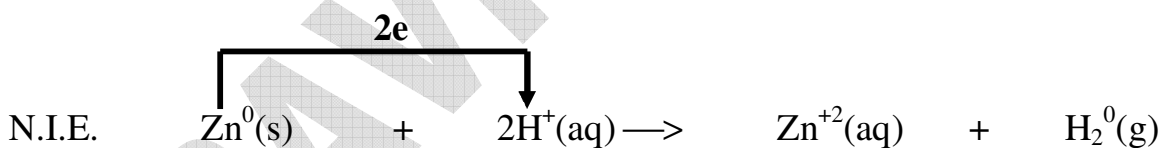
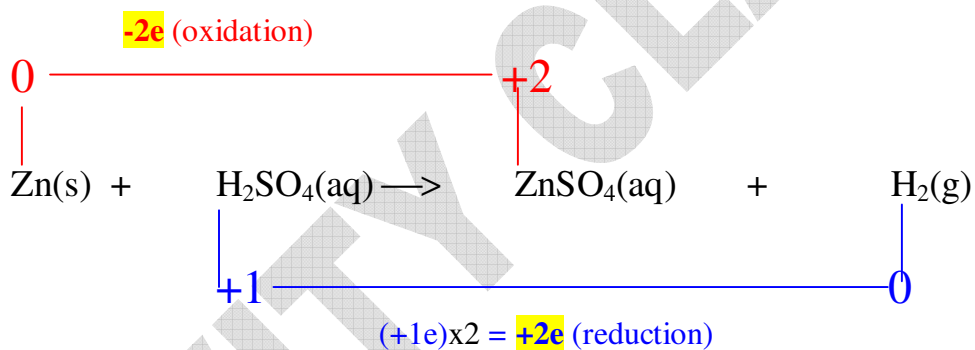
- Zn is more active than Cu
- Zn loses electrons easier than Cu
- Zn is a stronger reducing agent than Cu

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:

Zinc strip
Sulfuric acid
Bubbles of hydrogen on zinc strip



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

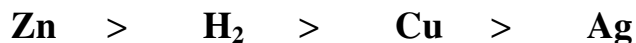
Since Zn replaces H^+ from its compound:

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

Summingup:

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 | VeryActive: React with water to produce H ₂ Ex: $\text{Na(s)} + \text{H}_2\text{O(l)} \longrightarrow \text{NaOH(aq)} + \text{H}_2\text{(g)}$ |
| Ba | |
| Ca | |
| Na | |
| Mg | ModeratelyActive: React with nonoxidizing acids to produce H ₂ gas Ex: $\text{Zn(s)} + \text{H}_2\text{SO}_4\text{(aq)} \longrightarrow \text{ZnSO}_4\text{(aq)} + \text{H}_2\text{(g)}$ |
| Al | |
| Zn | |
| Cr | |
| Fe | |
| Ni | |
| Sn | |
| Pb | |
| H ₂ | Relativelyinactive: Do not react with nonoxidizing acids to produce hydrogen gas. |
| Cu | |
| Ag | |
| Hg | |
| Au | |

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.

4. Combustion Reactions

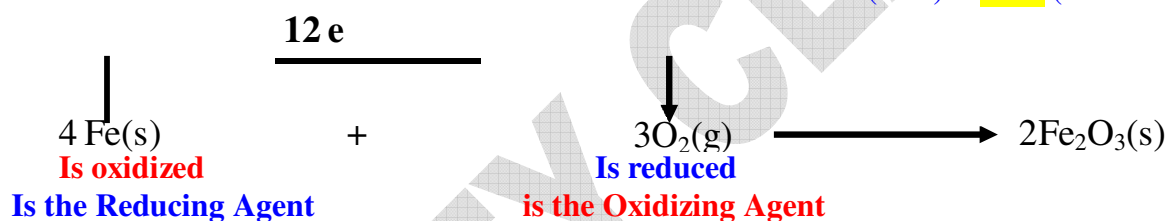
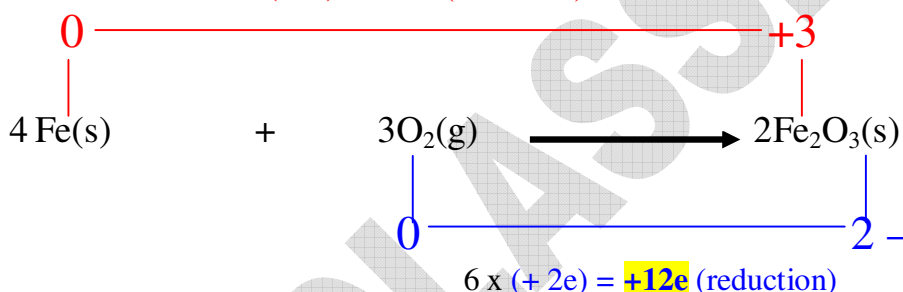
- Combustion reactions are those reactions in which a substance reacts with oxygen, with the rapid release of heat to produce a flame.

Example1:

Iron wool burns in air (reacts with oxygen) to produce iron (III) oxide.



$4 \times (-3e) = -12e$ (oxidation)



NOTE:

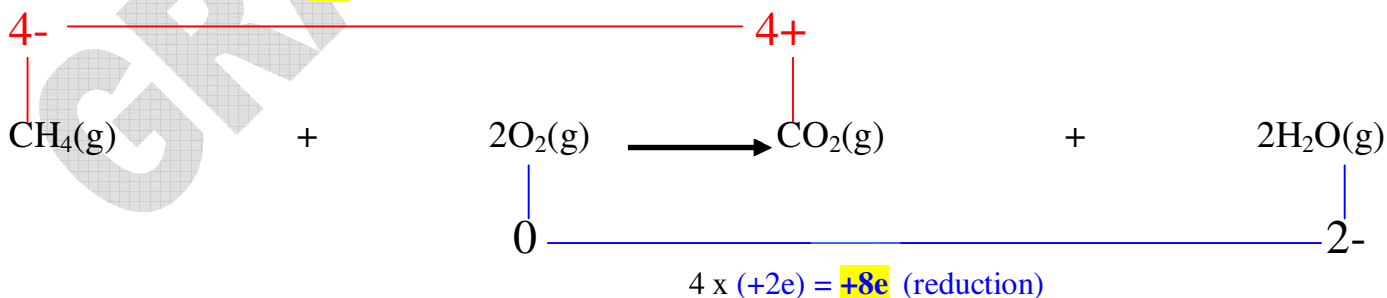
The above reaction can also be classified as a Combination Reaction.

Example2:

Methane gas (CH₄) burns in air (reacts with oxygen) to produce carbon dioxide and water vapor.



$-8e$ (oxidation)



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