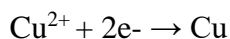


CHAPTER III : Redox Reactions

1. Definitions**1.1. An Oxidant (Oxidizing Agent)**

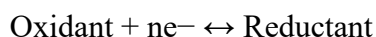
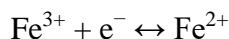
An **oxidant**, is a chemical species that **gains one or more electrons** in a reaction, causing another species to lose electrons. When an oxidant gains electrons, it gets **reduced**.

Example**1.2. A Reductant (Reducing Agent)**

A **reductor**, is a chemical species that is capable of **donating one or more electrons** in a reaction, When a reductant loses electrons, it gets **oxidized**.

Example**1.3. Redox Couple (Oxidant / Reductant)**

A **redox couple** consists of an **oxidant** and a **reductant** that correspond to each other in a half-reaction. The species **Ox** and **Red** form a **redox couple**, written as **Ox/Red**.

**Examples****Fe³⁺ / Fe²⁺ Couple**

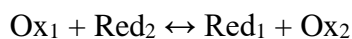
In this redox couple :

-**Fe³⁺** (ferric ion) is the **oxidant** (it can gain an electron to become Fe²⁺).

-**Fe²⁺** (ferrous ion) is the **reductant** (it can donate an electron to become Fe³⁺).

2. Redox Reaction

A redox reaction (oxidation-reduction reaction) involves **two redox couples**, where electrons are transferred between an **oxidant** from one redox couple and a **reductant** from another redox couple. The reaction can be written as :

**Example 1**

Note : In a **basic** medium, OH^- ions are used to balance **oxygen** and H_2O molecules to balance **hydrogen**, whereas in an **acidic** medium, H_2O molecules balance oxygen and H^+ ions balance hydrogen.

3. Oxidation Number (Oxidation State)

The **oxidation number (NO)** (also called the **oxidation state**) of an atom in a compound is the **number of electrons** an atom has gained or lost compared to its neutral (uncombined) state.

a. The oxidation number (NO) of an atom in its **free** or **elemental state** is always **zero**. This is because, there is no loss or gain of electrons in the neutral atom in this state.

Example

H, O, Cu, Co \rightarrow **NO = 0**

b. When two atoms of the **same element** combine to form a molecule, and no charge is present on the molecule (neutral combination), the **oxidation number** of each atom in the molecule is **zero**.

Example

$\text{F}_2, \text{Br}_2, \text{Cl}_2 \rightarrow$ **NO = 0**

c. The **oxidation number (NO)** of an atom in a **monoatomic ion** is equal to the **charge** of the ion.

Example

$\text{Cl}^- \rightarrow$ **NO (Cl) = -I** ; $\text{Fe}^{3+} \rightarrow$ **NO (Fe) = +III**.

d. In a **neutral molecule**, the **sum of the oxidation numbers** of all the atoms within the molecule is always **zero**.

Example

NH_3 : $\Sigma \text{NO} = \text{N.O(N)} + 3 \times \text{N.O(H)} = 0$

e. Oxygen is the most **electronegative** element after fluorine. In most compound, the oxidation number (NO) of oxygen is -2. However, there are exceptions depending on the specific bonding environment of oxygen.

Example

$\text{F}_2\text{O} \rightarrow$ **NO (F) = -I** and **NO (O) = +II**

$\text{H}_2\text{O}_2 \rightarrow$ **NO (H) = +I** and **NO(O) = -I**

f. The oxidation number (NO) of **hydrogen** varies depending on the compounds it forms.

Example

H_2 : **NO (H) = 0**.

Hydrogen in compounds with non-metals (H_2O , HCl) : NO (H) = +1.

Hydrides (NaH , KH , LiH) : NO (H) = -1 (hydrogen is bonded to metals that are more electropositive).

g. For a **polyatomic ion**, the **sum of the oxidation numbers (NO)** of all the atoms within the ion is equal to the **total charge** of the ion.

Example

Sulfate ion (SO_4^{2-}) : $\Sigma \text{NO} = \text{NO}(\text{S}) + 4 \times \text{NO}(\text{O}) = -2 \rightarrow \text{NO}(\text{S}) = +6$

4.