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

$Cu^{2+} + 2e- \rightarrow Cu$

1.2. A Reductant (Reducing Agent)

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

Example

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

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**.

Oxidant + ne− ↔ Reductant

Examples

Fe^{3+} / Fe^{2+} Couple

$Fe^{3+} + e^{-} \leftrightarrow Fe^{2+}$

In this redox couple :

-Fe^{3+} (ferric ion) is the **oxidant** (it can gain an electron to become Fe^{2+}).

-Fe^{2+} (ferrous ion) is the **reductant** (it can donate an electron to become Fe^{3+}).

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 :

$Ox_1 + Red_2 \leftrightarrow Red_1 + Ox_2$

Example 1

$Zn \rightarrow Zn^{2+} + 2e^{-}$ (**Oxidation Half-Reaction)**

$Cu^{2+} + 2e^{-} \rightarrow Cu$ (**Reduction Half-Reaction)**

$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ (**Redox Reaction)**

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 → **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

F_2, Br_2, Cl_2 → **NO = 0**

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

Example

Cl^- → **NO** (Cl) = **-I** ; Fe^{3+} → **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 : $\Sigma NO = N.O(N) + 3 \times N.O(H) = 0$

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

F2O → NO (F) = -I and NO (O) =+II

H2O2 → 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-}) : ΣNO = NO(**S**) + 4×NO(**O**) = -2 → NO(**S**) = +6

4.