UNIT 6 - ELECTROCHEMISTRY

Redox

A redox reaction is a reaction in which electron transfer takes place. In redox reactions, oxidation, and reduction will always occur simultaneously.

OXIDATION IS	REDUCTION IS
gain of oxygen	gain of hydrogen
loss of hydrogen	loss of oxygen
loss of electrons	gain of electrons
increase in oxidation state	decrease in oxidation state

<u>Oxidation</u>

1. Oxidation of an element:

$x+2(-2)=0$

$$\overset{0}{S}(s) + \overset{0}{O_2}(g) \longrightarrow \underset{+4}{\overset{2(-2)}{SO_2}}(g)$$

2. Oxidation of a compound:

$$\overset{-4\,(+1)4}{CH_4}(g) + 2O_2(g) \longrightarrow \overset{+4\;2(-2)}{CO_2}(g) + 2H_2O(l)$$

3. Oxidation using an indirect source of oxygen:

$$8K\overset{+5}{Cl}O_3(s) + \overset{-4}{C_{12}}H_{22}O_{11}(s) \longrightarrow 12\overset{+4}{C}O_2(g) + 11H_2O(l) + 8K\overset{-1}{Cl}(s)$$

$+1+x+3(-2)=0$

<u>Reduction</u>

1. Reduction to an element:

$$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$$

2. Reduction of one element by another:

$$2CuO(s) + C(s) \longrightarrow 2Cu(s) + CO_2(g)$$

<u>Redox</u>

There are two ways to find out whether a substance has been oxidized or reduced during a chemical reaction:

1) Electron transfer
2) Changes in the oxidation state

A balanced ionic equation can be constructed by balancing the number of electrons lost or gained in a redox reaction must be equal.

Half equations

$$2Na_{(s)} + Cl_{2\,(g)} \rightarrow 2NaCl_{(s)}$$

This reaction can be divided into two separate equations, one showing oxidation and the other reduction. These are known as half equations.

1) $Na \rightarrow Na^+ + e^-$

Each sodium atom loses an electron from its outer shell. Sodium atoms have been oxidized.

2) $Cl_2 + 2e^- \rightarrow 2Cl^-$

Each chlorine atom gains an electron to complete its outer shell. The chlorine atoms have been reduced.

Question: write two half-equations for the following reactions. For each half-equation state whether oxidation or reduction is occurring:

(a) $Cl_2 + 2I^- \longrightarrow I_2 + 2Cl^-$

$Cl_2 + 2e \longrightarrow 2Cl^-$ (red)

$2I^- \longrightarrow I_2 + 2e$ (ox)

(b) $2Mg + O_2 \longrightarrow 2MgO$

$2\times(Mg \longrightarrow Mg^{2+} + 2e)$

$O_2 + 4e^- \longrightarrow 2O^{2-}$

$2Mg \longrightarrow 2Mg^{2+} + \cancel{4e}$

$O_2 + \cancel{4e} \longrightarrow 2O^{2-}$

$2Mg + O_2 \longrightarrow 2MgO$

(c) $4Fe + 3O_2 \longrightarrow 2Fe_2O_3$

Question: zinc metal reacts with IO_3^- ions in an acidic solution. Construct a balanced ionic equation for this reaction, using the two half-equations below:

$$2IO_3^- + 12H^+ + 10e^- \longrightarrow I_2 + 6H_2O$$

$$(Zn \longrightarrow Zn^{2+} + 2e^-) \times 5$$

$$2IO_3^- + 12H^+ + 5Zn \longrightarrow I_2 + 6H_2O + 5Zn^{2+}$$

Oxidation Numbers

The oxidation number is the number of electrons that an individual atom within an ion or compound has either lost control of (+ve) or gained control of (-ve).

In simple ionic compounds, the non-metal is assigned a negative oxidation number and the metal is assigned a positive value.

In simple covalent compounds and complex ions, the most electronegative element is assigned a negative value, and the least electronegative element is assigned a positive value.

The oxidation number of an uncombined element is 0, for example, Na, Mg, K, O_2, and N_2.

Certain elements can usually be regarded as having fixed values in all of their compounds. Examples of these include:

- Group 1 elements are always +1
- Group 2 elements are always +2
- Group 3 elements are always +3
- Fluorine is always -1
- Oxygen is nearly always -2 unless combined with Flourine, or when it's in the form of a peroxide or superoxide.

The total oxidation number of all atoms in a neutral compound always adds up to 0.

∟ The total oxidation numbers of all the atoms in a complex ion always add up to the charge on the ion.

Oxidizing or Reducing Agent?

An **oxidizing agent** increases the oxidation number of another atom and gets reduced itself. An atom of the oxidizing agent decreases in oxidation numbers. The oxidizing agent gets reduced – it gains electrons.

A **reducing agent** decreases the oxidation number of another atom and gets oxidized. An atom of the reducing agent increases in oxidation number. The reducing agent gets oxidized – it loses electrons.