ST. LAWRENCE HIGH SCHOOL

A JESUIT CHRISTIAN MINORITY INSTITUTION

STUDY MATERIAL FOR CHEMISTRY (CLASS-11)

TOPIC- REDOX EQUILIBRIA

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Redox reaction is the combination of both oxidation and reduction reactions. When a redox reaction takes place, both oxidation and reduction reactions also occur simultaneously.

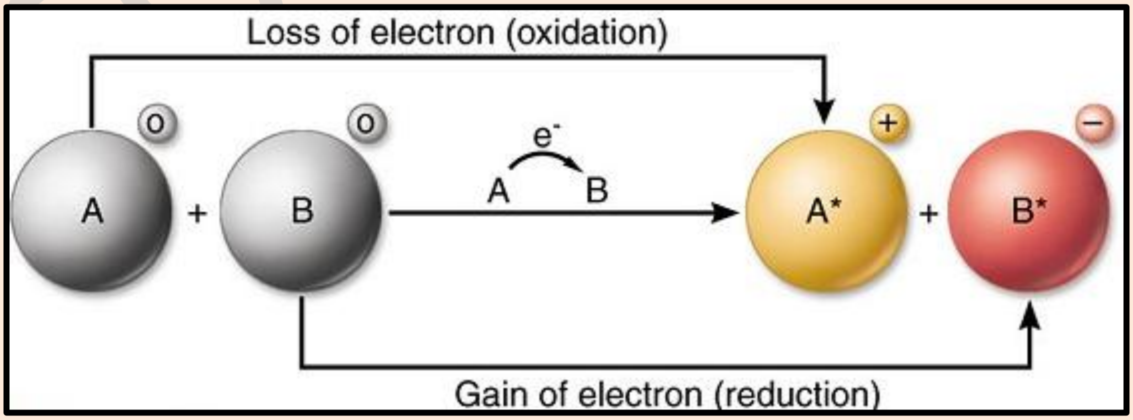
WHAT IS A REDOX REACTION?

Redox is a term used for the oxidation-reduction reaction. Redox reaction is a chemical reaction where a change in the oxidation state of atoms occurs. It involves electron transfer, i.e. if one chemical species gains electrons, then another chemical species gives or loses electrons. The species from which the electron is lost is said to be oxidized whereas the species to which the electron is added is said to be reduced. Example zinc displaces copper in an aqueous solution called copper sulfate.

TYPES OF REDOX REACTIONS

Redox reactions can be classified into four different categories:

- Combination reaction
- Decomposition reaction
- Displacement reaction
- Disproportionation reaction
- Comproportionation reaction



Redox Reaction

In the illustration provided above, it can be observed that the reactant an electron was removed from the reactant A, and this reactant is oxidized. Similarly, reactant B was handed an electron and was therefore reduced.

The loss of electrons and the corresponding increase in the oxidation state of a given reactant is called oxidation. The gain of electrons and the corresponding decrease in the oxidation state of a reactant is called reduction.

Electron-accepting species which tend to undergo a reduction in redox reactions are called oxidizing agents. An electron-donating species which tends to hand over electrons can be referred to as a reducing agent. These species tend to undergo oxidation. It can be noted that any redox reaction can be broken down into two half-reactions, namely the oxidation half-reaction and the reduction half-reaction.

When writing these half-reactions separately, each of them must be balanced in a way that all the electrons are accounted for.

1) Decomposition Reaction

This kind of reaction involves the breakdown of a compound into different compounds. Examples of these types of reactions are:

- 2NaH → 2Na + H2
- 2H2O → 2H2+O2
- Na2CO3 → Na2O + CO2

All the above reactions result in the breakdown of smaller chemical compounds in the form of AB → A + B

But, there is a special case that confirms that all the decomposition reactions are not redox reactions.

For example $CaCO_3$ (s) → CaO (s) + CO_2 (g)

2) Combination Reaction

These reactions are the opposite of decomposition reaction and hence involve the combination of two compounds to form a single compound in the form of A + B → AB.

For example:

- $H_2 + Cl_2 \rightarrow 2HCl$
- $C+O_2 \rightarrow CO_2$
- $4Fe+3O_2 \rightarrow 2Fe_2O_3$

3) Displacement Reaction

In this kind of reaction, an atom or an ion in a compound is replaced by an atom or an ion of another element. It can be represented in the form of X + YZ → XZ + Y. Further displacement reaction can be categorized into

- Metal displacement Reaction
- Non-metal displacement Reaction

A) Metal Displacement

In this type of reaction, a metal present in the compound is displaced by another metal. These types of reactions find their application in metallurgical processes where pure metals are obtained from their ores.

B) Non-Metal Displacement

In this type of reaction, we can find a hydrogen displacement and sometimes rarely occurring reactions involving oxygen displacement.

4) Disproportionation Reactions

The reactions in which single reactant is oxidized and reduced is known as Disproportionation reactions.

For example: $P_4 + 3NaOH + 3H_2O \rightarrow 3NaH_2PO_2 + PH_3$

Disproportionation, sometimes called **dismutation**, is a redox reaction in which one compound of intermediate oxidation state converts to two compounds, one of higher and one of lower oxidation states

5) Comproportionation Reactions

Comproportionation or **synproportionation** is a chemical reaction where two reactants, each containing the same element but with a different oxidation number, form a product in which the elements involved reach the same oxidation number. It is opposite to disproportionation.

For example: $IO_3^- + 5\, I^- + 6\, H^+ \rightarrow 3\, I_2 + 3\, H_2O$

EXAMPLES OF REDOX REACTIONS

A few examples of redox reactions along with their oxidation and reduction half-reactions are provided in this subsection.

Example 1: Reaction between Hydrogen and Fluorine

In the reaction between hydrogen and fluorine, the hydrogen is oxidized whereas the fluorine is reduced. The reaction can be written as follows.

$$H_2 + F_2 \rightarrow 2HF$$

The oxidation half-reaction is: **$H_2 \rightarrow 2H^+ + 2e^-$**

The reduction half-reaction is: **$F_2 + 2e^- \rightarrow 2F^-$**

The hydrogen and fluorine ions go on to combine in order to form hydrogen fluoride.

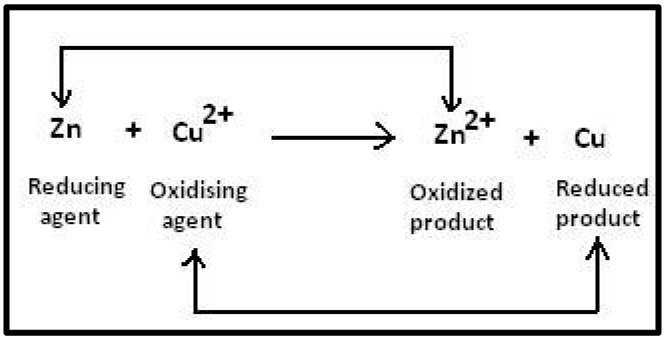
Example 2: Reaction between Zinc and Copper

This is a type of metal displacement reaction in which copper metal is obtained when zinc displaces the Cu^{2+}ion in the copper sulfate solution as shown in the reaction below.

$Zn\ (s) + CuSO_4\ (aq) \rightarrow ZnSO_4\ (aq) + Cu\ (s)$

The oxidation half-reaction can be written as: **$Zn \rightarrow Zn^{2+} + 2e^-$**

The reduction half-reaction can be written as: **$Cu^{2+} + 2e^- \rightarrow Cu$**



Thus, copper is displaced from the copper sulfate solution by zinc in a redox reaction.

Example 3: Reaction between Iron and Hydrogen Peroxide

Fe^{2+} is oxidized to Fe^{3+} by hydrogen peroxide when an acid is present. This reaction is provided below.

$2Fe^{2+} + H_2O_2 + 2H^+ \rightarrow 2Fe^{3+} + 2H_2O$

Oxidation half-reaction: **$Fe^{2+} \rightarrow Fe^{3+} + e^-$**

Reduction half-reaction: **$H_2O_2 + 2e^- \rightarrow 2\ OH^-$**

Thus, the hydroxide ion formed from the reduction of hydrogen peroxide combines with the proton donated by the acidic medium to form water.

OXIDATION AND REDUCTION REACTION

In order to understand redox reactions, let us first deal with oxidation and reduction reactions individually.

What is Oxidation Reaction?

Oxidation may be defined as loss of electrons from a substance, the other definition of oxidation reactions states that addition of oxygen or the more electronegative element or removal of hydrogen or the more electropositive element from a substance is called as an oxidation reaction.

Following are some examples of oxidation reactions:

$$2S(s) + O_2\,(g) \rightarrow SO_2\,(g)$$

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(l)$$

What is Reduction Reaction?

Like oxidation reactions, reduction reactions are defined as the gain of electrons. Any substance that gains electron during a chemical reaction gets reduced.

In other forms, the reduction reaction is stated as the addition of hydrogen or more electropositive element or removal of a more electronegative element or oxygen from a substance.

Below are some examples of reduction reactions:

- $2CH_2CH_2\,(g) + H_2\,(g) \rightarrow CH_3CH_3\,(g)$
- $2FeCl_3\,(aq) + H_2\,(g) \rightarrow 2FeCl_2\,(aq) + 2HCl\,(aq)$

Now if we closely examine the above reaction we would find that all the reactions above have both, reduction and oxidation reactions.

The reaction in which FeCl3 is getting reduced as electronegative element chlorine is being removed from it. While hydrogen is getting oxidized due to the addition of chlorine, an electronegative element, in the same reaction.

OXIDIZING AND REDUCING AGENTS

- The substance (atom, ion, and molecule) that gains electrons and is thereby reduced to a low valency state is called Oxidising agent.
- The substance that loses electrons and is thereby oxidised to a higher valency state is called a reducing agent.

Important Oxidizing Agents

- Molecules made up of electronegative elements. **e.g. O_2, O_3, and X_2 (halogens)**
- Compounds containing an element which is in the higher oxidized state. **eg. $KMnO_4$, $K_2Cv_2O_7$, HNO_3, $KClo_3$**
- Oxides of metals and non-metals **e.g.MgO, CuO, CrO_3, P_4O_{10}**
- Fluorine is the strongest oxidizing agent.

Important Reducing Agents

- All metals, **e.g** Na, Zn, Fe, Al
- A few non-metals **e.g** C, Hydrogen, S, P
- Hydracids, **e.g** HCl, HBr, HI, H_2S
- Few compounds containing an element in the lower oxidation state, **e.g**$FeCl_2$, $FeSo_4$, $SnCl_2$, Hg_2Cl_2
- Metallic hydrides, **e.g**NaH, LiH, CaH_2, etc..
- Organic compounds like HCOOH,

Lithium is the strongest reducing agent in the solution and Cesium is the strongest reducing agent in the absence of water. The substances which act as oxidizing as well as reducing agents are H_2O_2, SO_2, H_2SO_3, HNO_2, $NaNO_2$.

REDUCTION POTENTIAL OF A HALF-REACTION

Each of the half-reactions that make up a redox reaction has a standard electrode potential. This potential equals the voltage produced by an electrochemical cell in which the cathode reaction is the half-reaction considered, whereas the anode is a standard hydrogen electrode.

This voltage produced by the half-reactions is called their reduction potentials (denoted by $E^0{}_{red}$). The value of the reduction potential of a half-reaction is positive for the oxidizing agents that are stronger than H^+ and negative for the ones that are weaker.

Examples of the reduction potentials of some species are +2.866 V for F_2 and -0.763 V for Zn^{2+}.

IDENTIFICATION OF OXIDIZING AND REDUCING AGENTS

- If an element is in its higher possible oxidation state in a compound. It can function as an oxidising agent. eg. $KMnO_4$, $K_2Cr_2O_7$, HNO_3, H_2SO_4, $HClO_4$
- If an element is in its possible lower oxidation state in a compound, it can function as a reducing agent. eg. H_2S, $H_2C_2O_4$, $FeSO_4$, $SnCl_2$
- If a highly electronegative element is in its highest oxidation state, the compound will act as an oxidising agent.
- If a highly electronegative element is in its lowest oxidation state the compound acts as a reducing agent.

Example: Identify oxidizing agent and reducing agent in the reactions.

1) $2Na_2S_2O_3+I_2 \rightarrow Na_2S_4O_6+2NaI$

2) $2FeCl_3+H_2S \rightarrow 2FeCl_2+S+2HCl$

3)$AgCN+CN^- \rightarrow [Ag(CN)_2]^-$

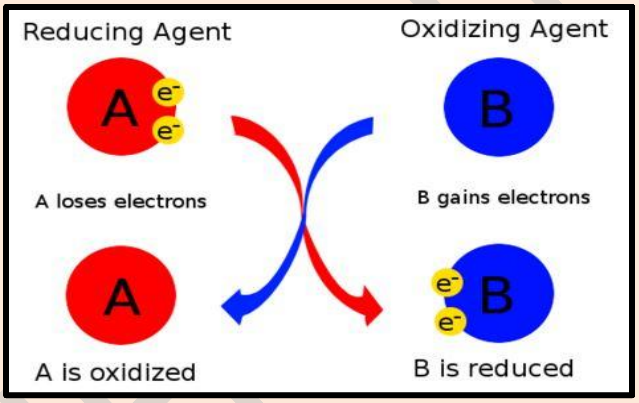
WHAT IS AN OXIDIZING AGENT?

An oxidizing agent (often referred to as an oxidizer or an oxidant) is a chemical species that tends to oxidize other substances, i.e. causes an increase in the oxidation state of the substance by making it lose electrons. Common examples of oxidizing agents include halogens (such as chlorine and fluorine), oxygen, and hydrogen peroxide (H_2O_2).

Definition

Oxidizing agents can be defined in two different ways:

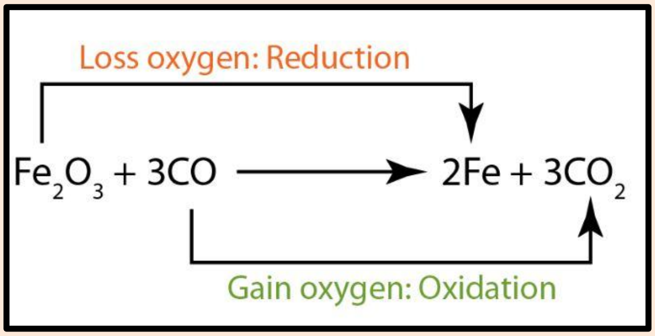
As an electron acceptor – They are chemical substances whose atoms remove at least one electron from another atom in a chemical reaction. As per this definition, oxidizing agents are the reactants that undergo reduction in redox reactions. An illustration detailing the electron-accepting properties of oxidizing agents is provided below.



Oxidizing Agents undergo Reduction by Gaining Electrons

Here, substance ‘A’ undergoes oxidation, resulting in an increase in its oxidation number. On the other hand, the oxidation state of substance ‘B’ becomes smaller (since it gains electrons by undergoing reduction).

As an atom-transferring substance – An oxidizing agent is a substance that transfers at least one electronegative atom to a chemical species in a chemical reaction. The transferred atom is typically an oxygen atom. Several combustion reactions and organic redox reactions involve the transfer of an electronegative atom between two reactants.



Oxidizing Agent that Transfers an Electronegative Atom

In the example illustrated above, the Fe_2O_3 molecule acts as an oxidizer by transferring an electronegative oxygen atom to the carbon monoxide molecule.

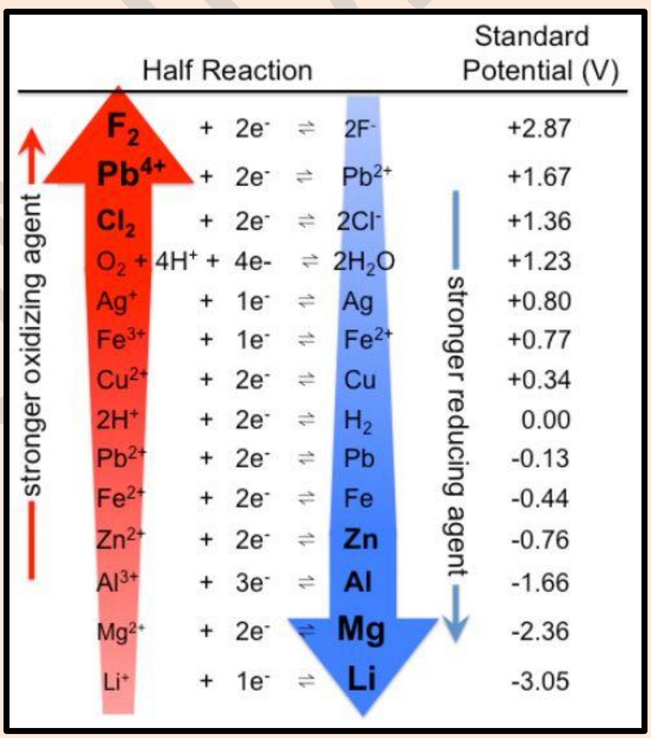
WHAT FACTORS AFFECT THE OXIDIZING POWER OF AN OXIDIZING AGENT?

Oxidizing agents normally exist in their highest possible oxidation states and, therefore, have a strong tendency to gain electrons and undergo reduction. Ions, Atoms, and molecules having a strong affinity towards electrons are considered to be good oxidizers. The stronger the electron affinity, the greater the oxidizing power.

Elemental fluorine is said to be the strongest elemental oxidizing agent. This is perhaps due to the fact that fluorine is the most electronegative element in the modern periodic table, and therefore exerts the strongest attractive force on electrons amongst all the elements. In fact, the oxidizing power of diatomic fluorine (F_2) is strong enough to cause metals such as asbestos and quartz (and even molecules, such as water) to burst into flames when exposed to it.

A few other examples of elemental oxidizing agents include diatomic oxygen (O_2), diatomic chlorine (Cl_2), and ozone (O_3). These oxidizers are the elemental forms of the second and the third most electronegative elements (oxygen and chlorine respectively), making them good electron acceptors.

The standard electrode potentialof a half-reaction in a redox process provides insight into the oxidizing power of the chemical substance. An illustration ranking some oxidizers in terms of their oxidizing powers is provided below.



Oxidizing Power of some Oxidizing Agents

Some compounds that exhibit large oxidation states can also be considered good oxidizing agents. Ionic examples include the permanganate ion, the chromate ion, and the dichromate ion. Acidic examples of good oxidizers include nitric acid, perchloric acid, and sulfuric acid. The electronegativity of the molecules increases with the increase in the oxidation states of the atoms, increasing their ability to oxidize other substances.

Examples of Oxidizing Agents

Halogens

The group 17 elements of the periodic table are collectively referred to as Halogens. They are said to have a strong ability to gain electrons, attributed to their high electronegativities when compared to elements from other groups. This implies that they have the ability to easily attract electrons towards their respective nuclei. Examples of the halogens that are good oxidizing agents include iodine, bromine, chlorine, and fluorine. Fluorine is said to be the strongest elemental oxidizing agent due to its highest electronegativity, as discussed earlier.

Oxygen

Oxygen is the element corresponding to the atomic number 8 and is denoted by the symbol 'O'. It belongs to the chalcogen group of the periodic table and is a highly reactive non- metal with good oxidizing properties. In general, metals tend to form metal oxides by reacting with atmospheric oxygen, due to the strong oxidizing power of oxygen. Oxygen is observed to be a part of a majority of combustion reactions.

Hydrogen Peroxide

Hydrogen peroxide is the chemical compound having formula H_2O_2. It appears to the human eye as a colourless liquid which has a greater viscosity than water. Hydrogen peroxide is the simplest compound having a peroxide functional group with an oxygen-oxygen single bond. It finds its uses as a weak oxidizing agent, disinfectant, and a bleaching agent.

Many other oxidizing agents are commonly used industrially as well as in the day to day lives of humans. Examples include household bleach (NaClO), Potassium Nitrate (KNO_3), and Sulfuric acid (H_2SO_4).

What are the Applications of Oxidizing Agents?

Oxidizing agents have numerous commercial and industrial applications. A few of these applications are listed below.

- Bleaching of fabrics.
- Purification of water.
- Combustion of fuel involves the use of an oxidizing agent.
- Storage of energy in batteries.
- Vulcanization of rubber (increasing the strength and the elasticity of rubber).

- Oxidizing agents are also vital to many biological processes such as metabolism and photosynthesis.

Many organisms make use of electron acceptors, or oxidizers, to collect energy from the redox reactions such as in the process of hydrolysis of glucose. To learn more about oxidizing agents and the part they play in oxidation-reduction reactions, register with BYJU'S and download the mobile application on your smartphone.

What are Reducing Agents?

The substance which loses electrons to the other substance and gets oxidized to the higher valency state is known as reducing agent.
A reducing agent is ***one of the reactants of an oxidation-reduction reaction which reduces the other reactant by giving out electrons to the reactant.*** If the reducing agent does not pass electrons to other substance in a reaction, then the reduction process cannot occur.

Characteristics of Reducing Agent

- The reducing agents give away electrons. The metals of the s-block in the periodic table are said to be good for reducing agents.
- These agents have an opposite effect to measuring the agents which tend to strengthen.
- The reducing agent after losing electrons gets oxidized and also causes the opposite reactant to get reduced by supplying electrons.
- All the good reducing agents have the atoms which have ***low electronegativity*** and a good ability of an atom or a molecule to attract the bonding electrons and the species having very small ionization energies.
- These usually serve as good reducing agents.
- All the ***oxidation and reduction reactions involve the transfer of electrons***.
- When some substance is oxidized, it is said to lose electrons and the substance which receives electrons is said to be reduced.
- If the substance has a strong tendency to lose electrons, then it is said to be strong reducing agent as it will reduce the other substances by giving electrons.

Strong vs Weak Reducing Agent

The stronger the reducing agent, the weaker is the corresponding oxidizing agent. Fluorine gas is known to be a strong oxidizing agent and whereas F- is said to be a weak reducing agent. We also know that – the weaker an acid then stronger is the conjugate base. In a similar way, the weaker the oxidizing agent then the stronger is the corresponding reducing agent as shown in the figure below.

$F_2(g) + 2e^- \longrightarrow 2F^-$

Strong oxidising agent — Weakest reducing agent

$Na^+ + e^- \longrightarrow Na$

Weakest oxidising agent — Strongest reducing agent

REDUCING AGENT EXAMPLE

Some of the common reducing agents include metals such as Na, Fe, Zn, Al and non-metals such as C, S, H_2. Some of the compounds and also the Hydracids such as HCl, HI, HBr, H_2S behave as good reducing agents. A brief explanation over some of the reducing agents are given below-

- **Lithium**– Lithium is a chemical element with atomic number 3 and a symbol Li. It appears like a soft and silvery-white metal and belongs to the alkali metal group of the periodic table. It is said to be a strong reducing agent when placed in solutions.
- **Iodides**– The salts of Iodides are said to be mild reducing agents. They react with oxygen to give out iodine. These also possess various antioxidant properties.
- **Reducing sugars**– Reducing sugars are those which behave in a similar as that of the reducing agents because of the free ketone group or a free aldehyde group present. All monosaccharides along with disaccharides, polysaccharides, oligosaccharides are said to be reducing sugars.

In order to learn how to find the oxidation number of an atom in a given compound, it is important to learn what oxidation numbers are. The oxidation number of an atom is a number that represents the **total number of electrons lost or gained** by it.

Calculating Oxidation Numbers

An oxidation number can be assigned to a given element or compound by following the following rules.

- Any free element has an oxidation number equal to zero.
- For monoatomic ions, the oxidation number always has the same value as the net charge corresponding to the ion.
- The hydrogen atom (H) exhibits an oxidation state of +1. However, when bonded with an element with less electronegativitythan it, it exhibits an oxidation number of -1.
- Oxygen has an oxidation of -2 in most of its compounds. However, in the case of peroxides, the oxidation number corresponding to oxygen is -1.
- All alkali metals (group 1 elements) have an oxidation state of +1 in their compounds.

- All alkaline earth metals (group 2 elements) exhibit an oxidation state of +2 in their compounds.
- In the compounds made up of two elements, a halogen (group 17 elements) have an oxidation number of -1 assigned to them.
- In the case of neutral compounds, the sum of all the oxidation numbers of the constituent atoms totals to zero.
- When polyatomic ions are considered, the sum of all the oxidation numbers of the atoms that constitute them equals the net charge of the polyatomic ion.

Thus, the oxidation number of an atom in a given compound can be calculated with the steps mentioned above.

Solved Examples

In order to help students understand how to find oxidation number, the oxidation states of each individual atom in some example compounds are determined below.

An illustration explaining how to find oxidation number of the sulfur atom in a sodium sulfate molecule can be found above.

NOTE: The second method for evaluating oxidation number is the consideration of molecular structure.

Hydrochloric Acid (HCl)

- As per the rules discussed above, the oxidation state of a group 17 element (halogen) in a diatomic molecule is -1. It is also discussed that hydrogen always exhibits an oxidation number of +1 unless it is paired with a less electronegative element.
- Since chlorine is more electronegative than hydrogen, an oxidation number of +1 can be assigned to the hydrogen atom in HCl.
- Therefore, the oxidation number of hydrogen is +1 and the oxidation of chlorine is -1 in HCl. These values can be verified by adding these oxidation numbers. Since the total is zero, which is the value of the oxidation number corresponding to a neutral molecule, the values are verified.

Carbon Dioxide (CO_2)

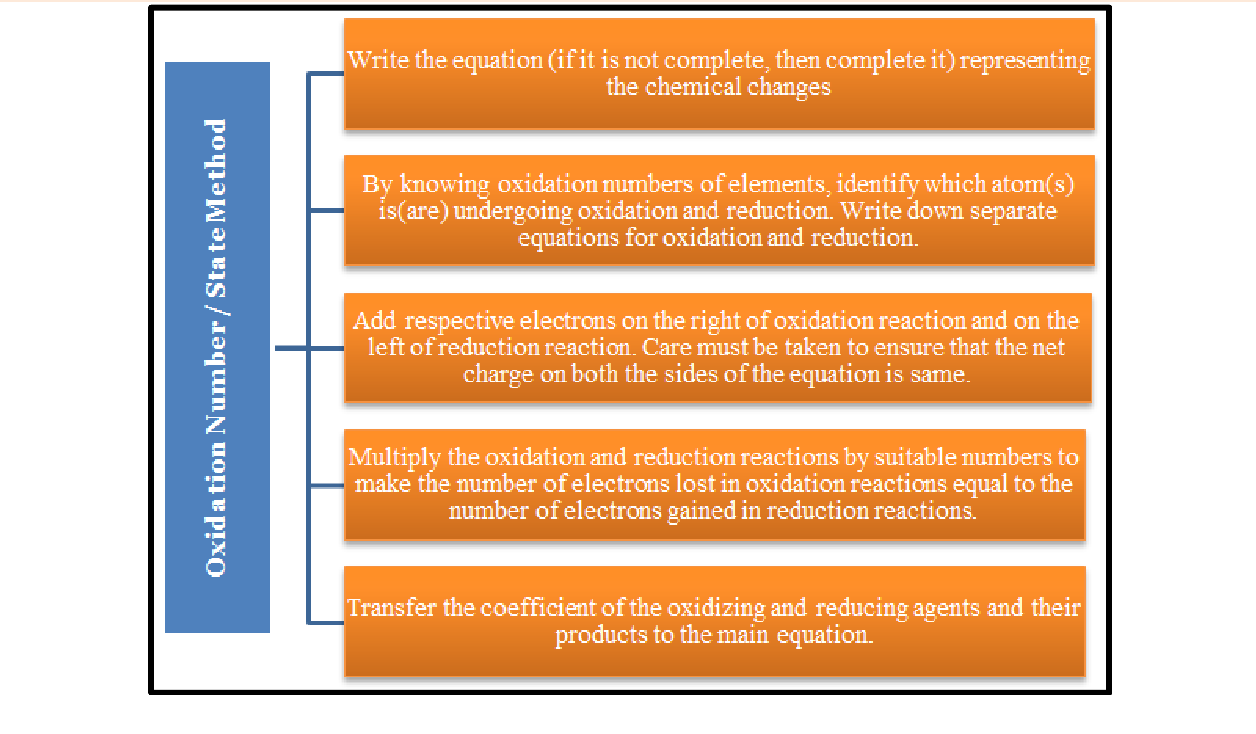
- According to the rules to calculate oxidation number, which can be found in the previous subsection, the oxidation number of oxygen in its compounds (excluding peroxides) is -2.
- Since there are two oxygen atoms in carbon dioxide, the total of the oxidation numbers corresponding to each oxygen is -4.
- Since the CO_2 molecule is neutral, the carbon atom must exhibit an oxidation state of +4 (the sum of all the oxidation numbers in a neutral molecule is zero).
- Therefore, the oxidation state of oxygen was found to be -2 and the oxidation number of carbon is +4 in a carbon dioxide molecule.

NOTE: Students can understand how to find oxidation number with the help of the solved examples provided above.

OXIDATION NUMBER / STATE METHOD FOR BALANCING REDOX REACTIONS

This method is based on the principle that the number of electrons lost in oxidation must be equal to the number of electrons gained in reduction. The steps to be followed are :

- Write the equation (if it is not complete, then complete it) representing the chemical changes.
- By knowing oxidation numbers of elements, identify which atom(s) is(are) undergoing oxidation and reduction. Write down separate equations for oxidation and reduction.
- Add respective electrons on the right of oxidation reaction and on the left of reduction reaction. Care must be taken to ensure that the net charge on both the sides of the equation is same.
- Multiply the oxidation and reduction reactions by suitable numbers to make the number of electrons lost in oxidation reactions equal to the number of electrons gained in reduction reactions.
- Transfer the coefficient of the oxidizing and reducing agents and their products to the main equation.
- By inspection, arrive at the co-efficient of the species not undergoing oxidation or reduction.



Question 1

Balance the equation:

$KMnO_4 + H_2SO_4 + FeSO_4 \rightarrow K_2SO_4 + MnSO_4 + Fe_2(SO_4)_3 + H_2O$

Solution:

$Mn^{+7} \rightarrow Mn^{+2}$; $Fe^{+2} \rightarrow Fe^{+3}$

$Mn^{+7} + 5e^- \rightarrow Mn^{+2}$ (1) $\times$ 2; $2Fe^{+2} \rightarrow Fe^{+3} + 2e^-$ (2) $\times$ 5

$2Mn^{+7} + 10e^- \rightarrow 2Mn^{+2}$; $10Fe^{+2} \rightarrow 5Fe^{+3} + 10e^-$.

Therefore coefficient of $KMnO_4$ is 2 that of $MnSO_4$ is 2.

Therefore $2KMnO_4 + 10FeSO_4 \rightarrow 2MnSO_4 + 5Fe_2(SO_4)_3$

Coefficient of K_2SO_4 should be 1.

$2KMnO_4 + 10\ FeSO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 5Fe_2\ (SO_4)_3$

To balance SO_4^{2-} so that left hand side must have $8H_2SO_4$, thus H_2O also becomes $8H_2O$.

$2KMnO_4 + 10\ FeSO_4 + 8H_2SO_4 \rightarrow K_2SO_4 + 2Mn\ (SO_4) + 5Fe_2\ (SO_4)_3 + 8H_2O$

Question 2:

Balance the equation:

$MnO_4^- + C_2O_4^{2-} + H^+ \rightarrow CO_2 + Mn^{+2} + H_2O$

Solution:

$Mn^{+7} \rightarrow Mn^{+2}$; $C_2 \rightarrow C^{+4}$

$5e^- + Mn^{+7} \rightarrow Mn^{+2}$ (1) ´2;

$C_2 \rightarrow 2C^{+4} + 2e^-$ (2) ´5

$10e^- + 2Mn^{+7} \rightarrow 2Mn^{+2}$; $5C_2 \rightarrow 10C^{+4} + 10e^-$

Therefore $2MnO_4^- + 5C_2O_4^{-2} \rightarrow 10CO_2^- + 2Mn^{+2}$

There must be $8H_2O$ on the right to balance O. Therefore there must be $10H^+$ to balance H.

$$2MnO_4^- + 5C_2O_4^{-2} + 16H^+ \rightarrow 10CO_2 + 2Mn^{+2} + 8H_2O$$

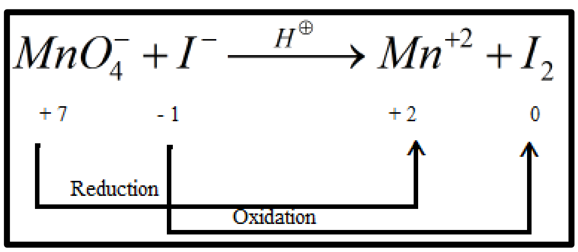
BALANCING REDOX REACTION BY ION ELECTRON METHOD (HALF REACTION METHOD)

Steps:

1. Divide the complete reaction into two half reaction, one representing oxidation and other representing reduction.
2. First balance other than ‘O’ and ‘H’ atoms.
3. In acidic or neutral medium balance oxygen atoms by adding H_2O molecule and balance H-atoms by adding H^+ ions.
4. In alkaline medium, oxygen atom are balanced by adding H_2O molecule and an equal number of OH^- ions are added on the opposite side, H^+ atoms still unbalanced add OH^-
5. Balance the charges by the addition of electrons.
6. Multiply with suitable integer such that the number of electrons gets cancelled.
7. Add both the half reactions, similar terms are subtracted and the final equation is written.

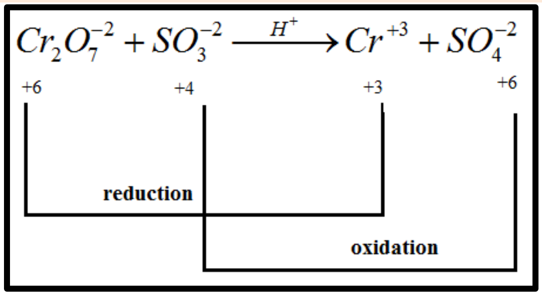
PROBLEMS ON BALANCING REDOX REACTIONS

Example 1: Balance the following redox reaction in the acid medium by ion-electron method



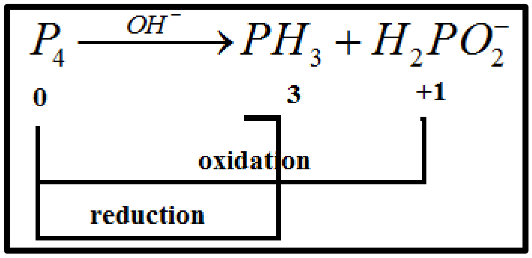
$2MnO_4^- + + 16H^+ + 10I^- \rightarrow 2Mn^{2+} + 5I_2 + 8H_2O$

Example 2: Balance the following redox reaction in the acid medium by ion-electron method



$Cr_2O_7^{-2} + 3SO_3^{-2} \rightarrow 2Cr^{+3} + 3SO_4^{-2} + 4H_2O + 8H^{+}$

<u>Example: 3</u> Balance the following reaction in basic medium by ion-electron method.



$$4P_4 + 12H_2O + 12OH^- \longrightarrow 12H_2PO_2^- + 4PH_3$$

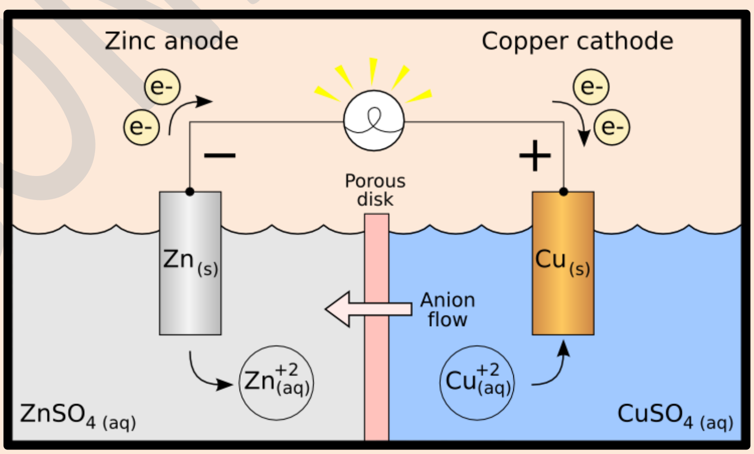
$$3P_4 + 3H_2O + 3OH^- \longrightarrow 3H_2PO_2^- + PH_3$$

Applications of Redox Reaction

Redox reactions have numerous industrial and everyday applications. A few of these applications of redox reactions are listed below.

Applications of Redox Reaction in Electrochemistry

The battery used for generating DC current uses redox reaction to produce electrical energy.



Applications of Redox Reaction

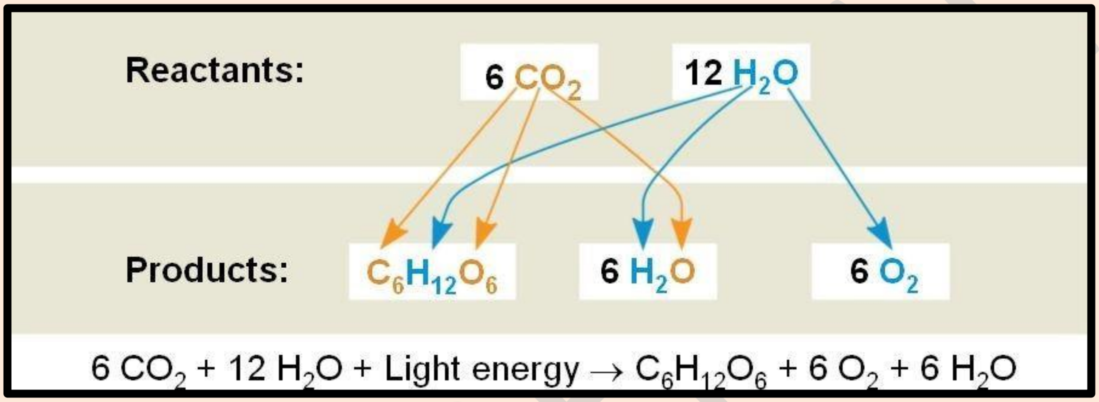
Batteries or electrochemical cellsused in our day-to-day life are also based on redox reactions. For example, storage cells are used in vehicles to supply all the electrical needs of the vehicles.

Redox Reaction in Combustion

Combustion is a type of oxidation-reduction reaction and hence it is a redox reaction. An explosion is a fast form of combustion and hence explosion can be treated as a redox reaction. Even the space shuttle uses redox reactions. The combination of ammonium perchlorate and powdered Aluminium inside the rocket boosters gives rise to an oxidation-reduction reaction.

APPLICATIONS IN PHOTOSYNTHESIS

Green plants convert water and carbon dioxide into carbohydrates and this process is defined as photosynthesis. The reaction is given as $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$



Applications of Redox Reaction in Photosynthesis

In the above reaction, we can see that carbon dioxide is reduced to carbohydrates while the water gets oxidized to oxygen and hence it is a redox reaction. The energy is provided by the sunlight for this reaction. This reaction is a source of food for animals and plants.

REAL LIFE USES OF REDOX REACTION

- Production of some important chemicals is also based on electrolysis which in turn is based on redox reactions. Many chemicals like caustic soda, chlorine, etc. are produced using redox reactions.
- Oxidation-Reduction reactions also find their application in sanitizing water and bleaching materials.
- The surfaces of many metals can be protected from corrosion by connecting them to sacrificial anodes which undergoes corrosion instead. A common example of this technique is the galvanization of steel.
- The industrial production of cleaning products involves the oxidation process.
- Nitric acid, a component of many fertilizers, is produced from the oxidation reaction of ammonia.
- Electroplating is a process that uses redox reactions to apply a thin coating of a material on an object. Electroplating is used in the production of gold-plated jewelry.
- Many metals are separated from their ores with the help of redox reactions. One such example is the smelting of metal sulfides in the presence of reducing agents.

The main source of oxidation is oxygen and therefore redox reaction or oxidation-reduction reactions are responsible for food spoilage.

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