CHEM 2101

Chapter - 4

Lecture Notes

Types of Chemical Reactions and solution Stoichiometry

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Chapter-4

Types of Chemical Reactions and solution Stoichiometry

(Most of the Problems in this chapter are already solved in the Lecture Class)

IMPORTANT TERMS

Solution

Solution is a homogeneous mixture of two or more substances.

Solute

Is present in lesser amount in the solution.

Solvent

Is present in greater amount in the solution.

Aqueous solution

Solutions formed by dissolving materials in the solvent water. These are called aqueous solutions.

Electrolyte

An electrolyte is a substance that when dissolved in water produces a solution that can conduct electricity.

Strong Electrolytes

Strong electrolytes are compounds, which produce a lot of ions in water solution. They are "completely dissociated" into ions in aqueous solution.

Strong electrolyte = solution that conducts electricity very well.

Weak Electrolytes

Weak electrolytes are materials that break into ions to a small extent. Weak electrolyte provides few ions in solution.

Weak electrolyte = solution that conducts electricity but not well.

Non electrolyte

Compounds that do not conduct electricity in aqueous solution are called nonelectrolytes.

These compounds do not give any ions at all in water solutions.

Cation

The positive ions are called cations.

Anion

The negative ions are called anions.

Concentration of solution

Tells the amount of solute present in a given amount of solvent or solution.

Strength of solution

The weight of solute in grams present in one liter of the solution.

Molarity

Molarity is defined as the number of gram moles of solute dissolved in 1 litre of solution.

Standard solution

A standard solution is a solution whose concentration is accurately known.

Dilution:

To save time and space in the laboratory, routinely used solutions are often purchased or prepared in concentrated form (called stock solution). Water is then added to achieve the molarity of desired particular solution. This process is called dilution.

Dilution with water does not alter the number of moles of solute present.

Precipitation reactions

When two solutions are mixed, an insoluble substance some times forms, that is a solid forms and separates from the solution. Such reaction is called a precipitation reaction, and the solid that forms is called precipitate.

Spectator ions

An ion that is presents during a reaction but does not undergo change; it appears on both sides of the reaction.

Arrehenius Concept of Acids and Base

An acid is a substance that produces H^+ ions when dissolved in water.

Base is a substance that produces OH^- ions when dissolved in water.

The Bronsted – Lowry Concept of acids and Bases

According to the Bronsted – Lowry Concept of acids and Bases:

An acid is a Proton Donor.

A base is a Proton Acceptor.

Neutralization Reactions

When acid and bases with equal amounts of hydrogen ion H^+ and hydroxide ions OH^- are mixed, the resulting solution is neutral.

Volumetric Analysis

Volumetric analysis is a technique for determining the amount of a certain substance by doing a titration.

Titration

A titration involves delivery (from a burette) of a measured volume of a solution of known concentration (the titrant) into a solution containing the substance being analysed (the analyte).

Indicator

Indicator is a substance added at the beginning of the titration that changes color at (or very near) the equivalence point.

End Point:

The point where the indicator actually changes color is called the end point of the titration.

Equivalence point:

The point in the titration where enough titrant has been added to react exactly with the analyte is called the equivalence point or the stoichiometric point.

Oxidation - Reduction reactions

The chemical reaction involving both oxidation and reduction is known as redox reaction.

Oxidation reaction:

Oxidation reaction is a process of addition of oxygen or removal of hydrogen.

Reduction reaction:

Reduction reaction is a process of removal of oxygen or addition of hydrogen.

Electronic concept of reduction and oxidation:

Oxidation is a process in which an atom taking part in a chemical reaction loses one or more electrons. The loss of electrons (LEO) results in the increase of positive charge.

Reduction is a process in which an atom taking part in a chemical reaction gains one or more electrons. The gain of electrons results in the decrease of positive charge. Gain of Electron Reduction (GER).

Oxidation number or Oxidation state

Oxidation number is defined as the residual charge which its atom has or appears to have when all other atoms from the molecule are removed as ions.

Atoms can have positive, zero or negative values of oxidation numbers depending upon their state of combination.

Reactions in Aqueous Solutions:

Many chemical reactions and virtually all biological processes take place in water.
So many do, that water is called the universal solvent.

General Properties of Aqueous Solutions:

1. A solution is a **homogeneous mixture** of two or more materials.

2. The **solvent is the material in the greater quantity** in the mixture. (Sometimes we think of it as the material that "does the dissolving".)

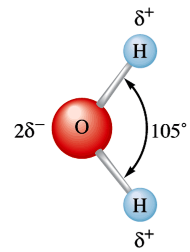
3. A **solute is a material in the smaller quantity** in the mixture.

4. Solutions may be homogeneous **mixtures of gases** (such as air), homogeneous **mixtures of solids** (such as an alloy of two or more metals like stainless steel), homogeneous **mixtures of liquids** (such as alcohol and water to make alcoholic beverages), or homogeneous mixtures of **solids or gases dissolved in a liquid**(such as salt water). **In this chapter we are dealing with solutions form by dissolving materials in the solvent water**. These are called **aqueous solutions.**

5. Melting points of solutions: Solutions are mixtures and therefore, the melting point of the mixture is lower than the melting point of the pure solvent.

6. Boiling points of solutions: Solutions are mixtures and therefore, the boiling Point of the mixture is higher than the boiling point of the pure solvent.

4.1 Water, the common Solvent



- **Water is the dissolving medium, or solvent.**

- **Water is "bent" or V-shaped.**
- **The O-H bonds are covalent.**
- **Water is a polar molecule.**
- **Hydration occurs when salts dissolve in water.**

How does water dissolve materials?

Water is a neutral molecule but it has a positive region (the H end of the molecule) and a negative region (the O end of the molecule). Neutral molecules with such positive and negative regions are called "polar molecules".

Because of its negative and positive "poles", water is a polar solvent.

Water dissolves many ionic compounds. The negative ends of water molecules attract and pull on the positive ions in the crystalline ionic lattice. The positive ends of water molecules pull negative ions in the crystalline ionic lattice.

Water dissolves many other polar molecules. If the solute is a very polar molecule, the negative ends of water molecules pull on the positive end of the polar molecule and the positive ends of water molecules pull on the negative end of the polar molecule and, if the "pull" is strong enough, the polar molecule can be pulled apart into ions in solution. The strong acid HCl is an example.

The ions in solution are surrounded by water molecules. The ions are said to be hydrated which is that each is surrounded by water molecules aligned according to the charge of the ion and the "pole" of the water molecule.

4.2 The Nature of Aqueous Solutions: Strong and Weak Electrolytes

Solution

Solution is a homogeneous mixture of two or more substances.

Solute

- **Dissolves in water (or other "solvent")**
- **Changes phase (if different from the solvent)**
- **Is present in lesser amount (if the same phase as the solvent)**

Solvent

- **Retains its phase (if different from the solute)**
- **Is present in greater amount (if the same phase as the solute)**

Aqueous solution

Solutions formed by dissolving materials in the solvent water. These are called aqueous solutions.

Electrolyte

An electrolyte is a substance that when dissolved in water produces a solution that can conduct electricity.

Strong Electrolytes

Strong electrolytes are compounds, which produce a lot of ions in water solution. They are "completely dissociated" into ions in aqueous solution.

Strong electrolyte = solution that conducts electricity very well.

Example:

Strong acids and Strong Base.

Weak Electrolytes

Weak electrolytes are materials that break into ions to a small extent. Weak electrolyte provides few ions in solution.

Weak electrolyte = solution that conducts electricity but not well.

Examples of weak electrolytes:

Weak acids
Weak bases like ammonia (NH_3)

Non electrolyte

Compounds that do not conduct electricity in aqueous solution are called nonelectrolytes.

These compounds do not give any ions at all in water solutions.

Non-electrolyes are materials that do not break into ions.

Examples: sugar, ethylene glycol (antifreeze), alcohol, etc.

Cation

The positive ions are called cations.

(because it was discovered that they were attracted to the cathode or the negative electrode).

Anion

The negative ions are called anions.

(because they are attracted to the anode).

Electrolytic Behavior

Strong Electrolytes are materials that **break up essentially 100% into ions in water solutions.**

Strong Acids: KNOW THESE 6 STRONG ACIDS.

HCl	in water is	$H^+(aq) + Cl^-(aq)$
HBr	in water is	$H^+(aq) + Br^-(aq)$
HI	in water is	$H^+(aq) + I^-(aq)$
HNO_3	in water is	$H^+(aq) + NO_3^-(aq)$
$HClO_4$	in water is	$H^+(aq) + ClO_4^-(aq$
H_2SO_4	in water is	$H^+(aq) + HSO_4^-(aq) + SO_4^{2-}(aq)$

Strong Bases: KNOW THESE.

The soluble hydroxides:

The IA family hydroxides such as NaOH and KOH
Plus only $Ba(OH)_2$ from other families

$NaOH(s)$	in water is	$Na^+(aq) + OH^-(aq)$
$KOH(s)$	in water is	$K^+(aq) + OH^-(aq)$

Weak electrolytes dissociate (ionize) only to a small extent in aqueous solution.

Example:

1. **The dissociation reaction for acetic acid (weak acid) in water can be written as follows:**

$$CH_3COOH(aq) \xLeftrightarrow{H_2O} H+(aq) + CH_3COO\text{-}(aq)$$

Acetic acid dissociates (ionizes) only to a slight extent in aqueous solutions is called a weak acid.

2. **When ammonia (weak base) dissolved in water, it reacts as follows:**

$$NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$$

Ammonia is called a weak base because the resulting solution is a weak electrolyte.

4.3 Composition of solutions

Chemical Reactions often take place when two solutions are mixed.

To perform stoichiometric calculations in such cases, we must know two things:

- **The nature of the reaction – which depends on the exact forms the chemicals take when dissolved.**
- **The amount of Chemicals present in the solution, usually expressed as concentration.**
- **The concentration of solution can be described in many different ways.**

Concentration of solution

Tells the amount of solute present in a given amount of solvent or solution.

Strength of solution

The weight of solute in grams present in one liter of the solution.

Concentration Units

There are many occasions when it is very important to know how much solute is present in a given amount of solvent. The amount of solute in a given amount of solvent is called the concentration.

There are lots of different units of concentration, which are useful for different situations. Ones that we will learn about at the appropriate times are wt/wt percent, wt/vol percent, vol/vol percent, ppm, ppb, mole fraction, etc.

The unit that is most important for the stoichiometry of solutions is called molarity.

Molarity

Molarity is defined as the number of gram moles of solute dissolved in 1 litre of solution.

$$\text{Molarity (M)} = \frac{\text{No. of moles of the solutes}}{\text{Volume of the solution in liters}}$$

In short we used to write

$$M = \frac{Mol}{V}$$

Moles of solute:

If we multiply the molarity of a solution by the volume (in liters) of a particular sample of the solution, we get the moles of solute present in that sample:

$$\text{Moles of solute} = \text{Liters of solution} \times \text{Molarity}$$

Standard solution

A standard solution is a solution whose concentration is accurately known.

Dilution:

To save time and space in the laboratory, routinely used solutions are often purchased or prepared in concentrated form (called stock solution). Water is then added to achieve the molarity of desired particular solution. This process is called dilution.

Dilution with water does not alter the number of moles of solute present.

Now we can do stoichiometry with solutions. All we have to do is add some new paths to our stoichiometry roadmap.

Recall that our old roadmap for stoichiometry was

$$\text{wt} \rightarrow \text{mol} \rightarrow \text{new mol} \rightarrow \text{new wt}.$$

Our expanded roadmap is

$$\begin{tikzcd}
\text{wt} \arrow[dr] & & & \text{new wt} \\
& \text{mol} \arrow[r] & \text{newmol} \arrow[ur] \arrow[dr] & \\
\text{vol} \arrow[ur] & & & \text{new vol}
\end{tikzcd}$$

This new roadmap offers all sorts of possibilities. We can start with either a weight or a volume and end up with either a weight or a volume (or even a new molarity).

Stoichiometry is used a lot in quantitative analyses.

4.4 Types of Chemical Reactions

There are three types of reaction in solution

1. **Precipitation reactions**
2. **Acid-base reactions**
3. **Oxidation-reduction reactions**

4.5 Precipitation Reactions

Simple Rules for the solubility of salts in water

1. **Most nitrate (NO_3^-) salts are soluble.**
2. **Most salts containing the alkali metal ions (Li+, Na+, K+, Cs+, Rb+) and the ammonium ion NH_4^+ are soluble.**
3. **Most chloride, bromide, and iodide salts are soluble. (Except salts containing the ions Ag^+, Pb^{2+} and Hg_2^{2+})**
4. **Most sulfate salts are soluble (Except $BaSO_4$, $PbSO_4$, $HgSO_4$, $CaSO_4$).**
5. **Most OH^- salts are only slightly soluble (NaOH, KOH are soluble. The compounds $Ba(OH)_2$, $Sr(OH)_2$ and $Ca(OH)_2$ are marginally soluble).**
6. **Most sulfide (S^{2-}), Carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), phosphate (PO_4^{3-}) salts are only slightly soluble.**

Precipitation reactions

- **When two solutions are mixed, an insoluble substance some times forms, that is a solid forms and separates from the solution. Such reaction is called a precipitation reaction, and the solid that forms is called precipitate.**

- **A precipitation reaction also can be called a double displacement reaction.**

- **When ionic compounds dissolve in water, the resulting solution contains the separated ions.**

Example:

$$AgNO3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO3(aq)$$

$$K2CrO4(aq) + Ba(NO3)2(aq) \rightarrow BaCrO4(s) + 2KNO3(aq)$$

4.6 Describing Reactions in solution

Three types of Equations are used to describe Reactions in Solution

- **The molecular equation stoichiometry but not necessarily the actual forms of the reactants and products in solution.**

- **The complete ionic equation represents as ions all reactants and products that are strong electrolyte.**

- **The net ionic equation includes only those solution components that undergo changes in the reaction. Spectator ions are not included.**

Spectator ions

An ion that is presents during a reaction but does not undergo change; it appears on both sides of the reaction.

Example:

For the following reaction, write the molecular equation, the complete ionic equation, and the net ionic equation.

Aqueous potassium chloride is added to aqueous silver nitrate to form a silver chloride precipitate plus aqueous potassium nitrate.

1. Molecular equation: (reactants and products as compounds)

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

2. Complete ionic equation: (all strong electrolytes shown as ions)

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

3. Net ionic equation: (show only components that actually react)

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Na^+ and NO_3^- are spectator ions.

4.7 Stoichiometry of Precipitation Reactions

Solving Stoichiometry Problems for Reactions in Solution

1. **Identify the species present in the combined solution, and determine what reaction occurs.**

2. **Write the balanced net ionic equation for the reaction.**

3. **Calculate the moles of reactants.**

4. **Determine which reactant is limiting.**

5. **Calculate the moles of product or product as required.**

6. **Convert to grams or other units as required.**

Species present → write the reaction → Balance net ionic equation → determine moles of reactants → Identify limiting reactant → determine moles of product → Check units of Products.

Now we can do stoichiometry with solutions. All we have to do is add some new paths to our stoichiometry roadmap.

Recall that our old roadmap for stoichiometry was

$$\text{wt} \rightarrow \text{mol} \rightarrow \text{new mol} \rightarrow \text{new wt.}$$

Our expanded roadmap is

$$\begin{tikzcd}
\text{wt} \arrow[dr] & & & \text{new wt} \\
& \text{mol} \arrow[r] & \text{newmol} \arrow[ur] \arrow[dr] & \\
\text{vol} \arrow[ur] & & & \text{new vol}
\end{tikzcd}$$

This new roadmap offers all sorts of possibilities. We can start with either a weight or a volume and end up with either a weight or a volume (or even a new molarity).

Stoichiometry is used a lot in quantitative analyses.

4.8 Acid – Base Reactions

Arrehenius Concept of Acids and Base:

According to the Arrehenius concept of acids and Bases:

An acid is a substance that produces H^+ ions when dissolved in water.

Base is a substance that produces OH^- ions when dissolved in water.

Example:

Acids produce H^+ in aqueous solutions

$$HCl \xrightarrow{\text{water}} H^+(aq) + Cl^-(aq)$$

Bases produce OH^- in aqueous solutions

$$NaOH \xrightarrow{\text{water}} Na^+(aq) + OH^-(aq)$$

The Bronsted – Lowry Concept of acids and Bases:

According to the Bronsted – Lowry Concept of acids and Bases:

An acid is a Proton Donor.

A base is a Proton Acceptor.

Example:

An Acid/ Base Reaction is (Bronsted/Lowry Definition) a transfer of H^+ ions

$$HCl(aq) + H_2O(aq) \rightarrow H_3O^+(l) + Cl^-(aq)$$

$$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq) + H_2O(l)$$

Know strong and weak acids

Example of strong acids: (HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_4).

Example of weak acids: (If not one of the 6 strong acids, assume it is weak.)

Know strong and weak bases:

Example of strong bases: IA Alkali hydroxides and $Ba(OH)_2$

Example of weak bases: Ammonia and organic amines

Properties of Acids and Bases:

Acid Properties:

1. **Sour taste**

2. **Acids change litmus from blue to red**

3. **Acids react with the more reactive metals to yield H_2 gas.**

   Example:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

4. **Acids react with carbonates and bicarbonates (hydrogen carbonates) to produce carbon dioxide.**

5. **Water solutions of acids are electrolytes:**

   Strong acids = strong electrolytes
   Weak acids = weak electrolytes

Base Properties:

1. **Bitter taste**

2. **Slippery feel**

3. **Bases change litmus from red to blue**

4. **Water solutions of bases are electrolytes:**

   Strong bases = strong electrolytes.
   Weak bases = weak electrolytes

Acid/Base Reactions:

An Acid/ Base Reaction is (Bronsted/Lowry Definition) a transfer of H^+ ions

An acid + a base (when the base is a hydroxide) $\rightarrow$ a salt + water

$$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O(l)$$

An acid + a base (ammonia type base) $\rightarrow$ a salt dissolved in water

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$$

Neutralization Reactions

When acid and bases with equal amounts of hydrogen ion H^+ and hydroxide ions OH^- are mixed, the resulting solution is neutral.

$$\underset{\text{base}}{NaOH\ (aq)} + \underset{\text{acid}}{HCl(aq)} \rightarrow \underset{\text{salt}}{NaCl\ (aq)} + \underset{\text{water}}{H_2O}$$

$$\underset{\text{base}}{Ca(OH)_2(aq)} + \underset{\text{acid}}{2HCl(aq)} \rightarrow \underset{\text{salt}}{CaCl_2(aq)} + \underset{\text{water}}{2H_2O}$$

- **Acid-base reactions involve a proton transfer.**
- **The acid donates a proton to the base.**
- **Acid-base reactions are also known as neutralization reactions.**
- **Acid + Base $\rightarrow$ Salt + water**
- **$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ is the most general neutralization reaction.**
- **Equivalence point is the point at which the moles of H^+ are equal to the moles of OH^-.**
- **An indicator is used to show the equivalence point during a titration.**
- **A titration involves the progressive addition of one reactant from burette (usually the Base), to a known volume of the other reactant in a conical flask (usually the acid).**

Neutralization

H_3O^+ and OH^- combine to produce water

$$\underset{\text{from acid}}{H_3O^+} + \underset{\text{from base}}{OH^-} \rightarrow \underset{\text{neutral}}{2H_2O}$$

Net ionic equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Ionic Equations for Neutralization

1. Molecular Equation:

   $$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

2. Complete Ionic equation:

   $$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

3. Net ionic equation:

   $$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Acid – Base Titrations

Volumetric Analysis:

Volumetric analysis is a technique for determining the amount of a certain substance by doing a titration.

Titration:

A titration involves delivery (from a burette) of a measured volume of a solution of known concentration (the titrant) into a solution containing the substance being analysed (the analyte).

Indicator:

Indicator is a substance added at the beginning of the titration that changes color at (or very near) the equivalence point.

End Point:

The point where the indicator actually changes color is called the end point of the titration.

Equivalence point:

The point in the titration where enough titrant has been added to react exactly with the analyte is called the equivalence point or the stoichiometric point.

Requirements for a successful titration

- **The exact reaction between titrant and analyte must be known.**
- **The stoichiometric (equivalence) point must be marked accurately.**
- **The volume of titrant required to reach the stoichiometric point must be known accurately.**

Performing Calculations for Acid – Base Reaction

1. **List the species present in the combined solution before any reaction will occur.**

2. **Write the balanced net ionic equation for this reaction.**

3. **Calculate the moles of reactants. For reactions in solution, use the volumes of the original solutions and their molarities.**

4. **Determine the limiting reactant where appropriate.**

5. **Calculate the moles of the required reactant or product.**

6. **Convert to grams or volume (of solution), as required.**

4.9 Oxidation - Reduction reactions:

The chemical reaction involving both oxidation and reduction is known as redox reaction.

Oxidation reaction:

Oxidation reaction is a process of addition of oxygen or removal of hydrogen.

Reduction reaction:

Reduction reaction is a process of removal of oxygen or addition of hydrogen.

Electronic concept of reduction and oxidation:

According to electronic concept, Oxidation is a process in which an atom taking part in a chemical reaction loses one or more electrons. The loss of electrons (LEO) results in the increase of positive charge.

For (e.g.,)

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$
$$Cu \rightarrow Cu^{2+} + 2e^-$$

The species which undergo loss of electrons during the reactions are called reducing agents. So Fe^{2+}, Cu are reducing agent (or) reductant.

Reduction is a process in which an atom taking part in a chemical reaction gains one or more electrons. The gain of electrons results in the decrease of positive charge. Gain of Electron Reduction (GER).

For (e.g)

$$Fe^{3+} + e^- \rightarrow Fe^{2+}$$
$$Zn^{2+} + 2e^- \rightarrow Zn$$

The species which undergo gain of electrons during the reactions are called oxidizing agents or oxidants. In the above reaction Fe^{3+}, Zn^{2+} are oxidizing agents.

Oxidation number or Oxidation state:

Oxidation number is defined as the residual charge which its atom has or appears to have when all other atoms from the molecule are removed as ions.

Atoms can have positive, zero or negative values of oxidation numbers depending upon their state of combination.

General Rules for assigning Oxidation Number to an atom:

The following rules are employed for determining oxidation number of atoms.

1. **The oxidation number of the element in the free (or) elementary state is always zero.**

   Example: Na*(s)*, O_2*(g)*, O_3*(g)*, Hg*(l)*

   Oxidation number of Helium in He= 0
   Oxidation number of chlorine in Cl_2 = 0

2. **The oxidation number of the element in monoatomic ion is equal to the charge on the ion.**

   Example: Na^+, Cl^-

3. **The oxidation number of Fluorine is always –1 in all its compounds.**

   Example: HF and PF_3

   The oxidation number of of Fluorine in HF and PF = -1.

4. **Hydrogen is assigned oxidation number +1 in all its compounds except in metal hydrides.**

   In covalent compounds Hydrogen is +1 oxidation state.
   Example: H_2O, HCl, NH_3.

   In metal hydrides the oxidation number of hydrogen is –1.

   Example: NaH, MgH_2, CaH_2, LiH etc.,

5. **Oxygen is assigned oxidation number –2 in most of its compounds, however in peroxides like H_2O_2, BaO_2, Na_2O_2, etc its oxidation number is –1. Similarly the exception also occurs in compounds of fluorine and oxygen like OF2 and O2F2 in which the oxidation number of oxygen is +2 and +1 respectively.**

6. The oxidation numbers of all the atoms in neutral molecule is zero. In case of polyatomic ion the sum of oxidation numbers of all its atoms is equal to the charge on the ion.

7. In binary compounds of non-metals, the more electronegative atom has negative oxidation number.

   Example: Oxidation number of K in KI is +1 but oxidation number of I is –1.

8. In binary compounds of non-metals, the more electronegative atom has positive oxidation number.

   Example: Oxidation number of Cl in ClF_3 is positive (+3) while that in ICl is negative (-1).

4.10. Balancing Oxidation – Reduction Reaction:

Oxidation – Reduction reactions in aqueous solutions are often complicated. Therefore it is difficult to balance their equations by simple inspection. In this section we will discuss a special technique for balancing the equations of redox reactions that occur in aqueous solutions. It is called the half-reaction method.

The Half – Reaction Method for Balancing Equations for oxidation – Reduction Reactions occurring in Acidic Solution:

1. **Find out the oxidation states of elements present in the reactants and products.**

2. **Write separate equations for the oxidation and reduction half reactions.**

3. **For each half reaction,**

   a. **Balance all the elements except hydrogen and oxygen.**
   b. **Balance oxygen using H_2O.**
   c. **Balance hydrogen using H^+.**
   d. **Balance the charge using electrons.**

4. **If necessary, multiply one or both balanced half reactions by an integer to equalize the number of electrons transferred in the two half reactions.**

5. **Add the half reactions, and cancel identical species.**

6. **Check that the elements and charges are balanced.**

The half-reaction method for balancing Equations for Oxidation – Reduction Reactions Occurring in Basic Solution:

1. **Use the half-reaction method as specified for acidic solutions to obtain the final balanced equation as if H+ ions were present.**

2. **To both sides of the equation obtained above, add a number of OH^- ions that is equal to the number of H^+ ions. (We want to eliminate H^+ by forming H_2O).**

3. **Form H_2O on the side containing both H^+ and OH^- ions, and eliminate the number of H_2O molecules that appear on both sides of the equation.**

4. **Check that elements and charges are balanced.**