I N D E X

IIT JAM PHYSICAL – THEORY BOOK

CHAPTER NO.	CHAPTER NAME	PAGE NO.
1	*MOLE CONCEPT*	*1*
2	*REDOX REACTIONS*	*10*
3	*ATOMIC STRUCTURE*	*17*
4	*GASEOUS STATE*	*34*
5	*LIQUID STATE*	*39*
6	THE SOLID STATE	*48*
7	*THERMODYNAMICS*	*57*
8	*CHEMICAL & PHASE EQUILIBRIA*	*75*
9	*ELECTROCHEMISTRY*	*98*
10	*CHEMICAL KINETICS*	*106*
11	*ADSORPTION*	*119*
12	*MOLECULAR SPECTROSCOPY*	*124*

CHAPTER 1
MOLE CONCEPT

MOLE CONCEPT

One mole is the amount of a substance that contains as many atoms/molecules/ particles, electrons, ions, etc as there are atoms in exactly 12 g (or 0.012 kg) of the ^{12}C isotope.

The mass of a one 12_C atom was determined by a mass spectrometer and it is equal to 1.992648×10^{-23} g. and one mole of carbon has weight 12 g,

So the number of atoms present in one mole of 12_C can be calculated as

∴ Number of atoms in 12 g 12_C is

$$= \frac{\text{Weight of 1 mole of } 12_C \text{ in g}}{\text{Weight of 1 Carbon atom}} = \frac{12 \text{ g mol}^{-1}}{1.992648 \times 10^{-23} \text{ g atom}^{-1}}$$

Number of atoms in 12 g 12_C is equal to $= 6.022 \times 10^{23}$ atoms mol^{-1}

That is 6.022×10^{23} number of entities (atoms/molecules/ particles, electrons, ions) are present in 1 mol of substance. This no 6.022×10^{23} is known as

Avogadro constant (N_A)

That is, 1 mol of hydrogen atom $= 6.022 \times 10^{23}$ atoms

1 mol of hydrogen molecule $= 6.022 \times 10^{23}$ molecules

1 mol of water $= 6.022 \times 10^{23}$ water molecules

1 mol of sodium ions $= 6.022 \times 10^{23}$ sodium ions

1 mol of electrons $= 6.022 \times 10^{23}$ electrons

1 mol of sodium chloride $= 6.022 \times 10^{23}$ formula units of sodium chloride

Molar Mass: The mass of one mole of a substance in grams is called its molar mass.

That is

Molar mass of water $= 18.02 \text{ g mol}^{-1}$

Molar mass of sodium chloride $= 58.5 \text{ g mol}^{-1}$

Mole in Terms of Volume

Mole is related to the volume of the gaseous substance.

Volume of one mole of any substance is called its molar volume. The molar volume of solids and liquids can be easily calculated if we know the molar mass and density at any given temperature and pressure because these do not change much with temperature and pressure.

However, the molar volumes of gases change considerably with temperature and pressure. It has been observed that one mole of an ideal gas (i.e., 6.022×10^{23} molecules) occupies 22.4 litres at N.T.P. (0°C and 1 atm pressure).

For example

1 mole of hydrogen gas at N.T.P. $= 22.4$ litres

1 mole of CO_2 gas at N.T.P. $= 22.4$ litres

This provides us with a method to determine molecular masses (weights) of gases from experimentally determined density as

Molecular mass $=$ Molar volume $\times$ Density

Some relations related to mole are given below:

i. One mole of atoms $= 6.022 \times 10^{23}$ atoms $=$ Gram atomic mass of the element

ii. Moles of an element $= \frac{\text{Mass of element}}{\text{Atomic mass}}$

iii. Mass of one atom $= \frac{\text{Atomic mass}}{6.022 \times 10^{23}}$

iv. One mole of molecules $= 6.022 \times 10^{23}$ molecules $=$ Gram molecular mass of the substance

v. Moles of a compound $= \frac{\text{Mass of compound}}{\text{Molecular mass}}$

vi. Mass of one molecule $= \frac{\text{Molecular mass}}{6.022 \times 10^{23}}$

vii. Volume occupied by 1 mole of a gas at N.T.P.= 22.4 L

viii. Mass of substance = moles of substance × Molar mass

ix. No of moles of substance $= \frac{\text{Volume at N.T.P. in Liters}}{22.4 \text{ L}}$

x. Volume at N.T.P. in Liters $= 22.4 \text{ L} \times$ No of moles of substance

xi. No of particles $= 6.022 \times 10^{23} \times$ No of moles of substance

xii. No of moles of substance $= \frac{\text{No of particles}}{6.022 \times 10^{23}}$

EXAMPLES

1. **Calculate the mass of an atom of silver (atomic mass = 108).**

   Answer:

   Mass of 6.022×10^{23} atoms of silver = 108 g

   Mass of 1 atom of silver = $\frac{108 \text{ g}}{6.022 \times 10^{23}}$

   = 1.793×10^{-22} **g**

2. **Calculate the mass of 1 molecule of water**

Answer:

6.022×10^{23} molecules of H_2O = Gram molecular mass = 18 g

Mass of 1 molecule of $H_2O = \frac{18}{6.022 \times 10^{23}}$

$= 2.99 \times 10^{-23}$ g

3. **Calculate the number of molecules and number of atoms present in 11.2 litres of oxygen (O_2) at N.T.P**

Answer:

We know that one mole of O_2 at N.T.P. occupies 22.4 litres.

11.2 litres of O_2 at N.T.P. $= \frac{1}{22.4} \times 11.2 = 0.5$ mole

Now, 1 mole of O_2 contains = 6.022×10^{23} molecules

0.5 mole of O_2 contains = $6.022 \times 10^{23} \times 0.5 = 3.01 \times 10^{23}$ molecules

1 molecule of oxygen = $2 \times 3.01 \times 10^{23} = 6.022 \times 10^{23}$ atoms

4. **The mass of 94.5 mL of a gas at S.T.P. is found to be 0.2231 g. Calculate its molecular mass.**

Answer:

Density of gas $= \frac{\text{Mass}}{\text{Volume}}$

$= \frac{0.2231 \text{ g}}{94.5 \text{ mL}}$

Molar volume $= 22.4$ L $= 22.4 \times 1000$ mL

Molecular mass = Molar volume × Density

Molecular mass $= \frac{22.4 \times 1000 \text{ mL} \times 0.2231 \text{ g}}{94.5 \text{ mL}}$

Molecular mass $= 52.9$ g

5. **Calculate the number of moles in 65.6 µg of carbon**

Answer:

Atomic mass of C =1 mole

12 g C= 1 mole

So 65.6×10^{-6} g = x moles

$\therefore$ x = moles of C $= \frac{\text{Mass}}{\text{Atomic mass}}$

$\therefore$ moles of C $= \frac{65.6 \times 10^{-6} \text{ g}}{12 \text{ g}}$

$\therefore$ moles of C $= 5.47 \times 10^{-6}$ moles

6. **Calculate the number of molecules in a drop of water weighing 0.05 g (H = 1, O = 16)**

Answer:

1 mole water $= 18$ g $= 6.022 \times 10^{23}$ molecules

Now 0.05 g =? molecules

no of molecules in 0.05 g water $= \frac{0.05 \text{ g} \times 6.022 \times 10^{23} \text{ molecules}}{18 \text{ g}}$

no of molecules in 0.05 g water $= 1.672 \times 10^{21}$ molecules

7. **Calculate the actual volume of one molecule of water**

Answer:

Density of water $= 1$ g mL^{-1}

$\therefore$ 1 g of $H_2O = 1$ mL

1 mole of water $= 18$ g of $H_2O = 18$ mL of water

$\therefore 6.022 \times 10^{23}$ molecules of water occupy volume $= 18$ mL

So

1 molecule of water occupy volume $= \frac{1 \text{ molecule} \times 18 \text{ mL}}{6.022 \times 10^{23} \text{ molecule}}$

1 molecule of water occupy volume $= 2.99 \times 10^{-23}$ mL

8. **Calculate the radius of a water molecule assuming to be spherical (density of water $= 1$ g mL^{-1})**

Answer:

Since water molecule is assumed to be spherical, its volume is equal to $4/3\ \pi r^3$

where r is the radius of water molecule.

Density of water $= 1$ g mL^{-1}

$\therefore$ 1 g of $H_2O = 1$ mL

1 mole of water $= 18$ g of $H_2O =$ 18 mL of water $\therefore 6.022 \times 10^{23}$ molecules of water occupy volume $= 18$ mL

So

1 molecule of water occupy volume $= \frac{1 \text{ molecule} \times 18 \text{ mL}}{6.022 \times 10^{23} \text{ molecule}}$

1 molecule of water occupy volume $= 2.99 \times 10^{-23}$ mL

$\therefore \frac{4}{3}\pi r^3 = 2.99 \times 10^{-23}$ mL

$r^3 = \frac{2.99 \times 10^{-23} \text{ mL} \times 3}{4 \times 3.14}$

$r = \left(\frac{2.99 \times 10^{-23} \text{ mL} \times 3}{4 \times 3.14}\right)^{1/3}$

$r = 1.925 \times 10^{-8}$ cm

$r = 1.925\ A^0$

Redox Reactions:

Oxidation and Reduction reactions

Oxidation reaction: Oxidation reaction can be identify by at least one of the following changes

i. It is the addition of oxygen to an element or a compound.

ii. Oxidation also means removal of hydrogen.
iii. Oxidation also means addition of other electronegative element.
iv. Oxidation also means removal of electropositive element

Now let's look at one by one change

i. Oxidation is the addition of oxygen to an element or a compound.

e.g. 2 Mg (s) + O_2 (g) → 2 MgO (s)

Here oxidation state of Mg changes from 0 to +2 and oxidation state of oxygen changes from 0 to -2

Oxidation state of Mg is increase by two; oxidation of Mg takes place

Oxidation state of oxygen is decrease by two; Reduction of Oxygen takes place

That is **Where there is oxidation, there is always reduction**

e.g. $S(s) + O_2(g) \rightarrow SO_2(g)$

Here oxidation state of S changes from 0 to + 4 and oxidation state of oxygen changes from 0 to -2

Oxidation state of S is increase by four; oxidation of S takes place

Oxidation state of oxygen is decrease by two; Reduction of Oxygen takes place

ii. Oxidation also means removal of hydrogen.

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

In this reaction hydrogen is replaced by oxygen.

$2H_2S(g) + O_2(g) \rightarrow 2S(s) + 2H_2O(l)$

In this reaction removal of hydrogen takes place that is oxidation takes place

iii. Oxidation also means addition of other electronegative element.

e.g.

$Mg(s) + F_2(g) \rightarrow MgF_2(s)$

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

$Mg(s) + S(s) \rightarrow MgS(s)$

In these reactions oxidation of magnesium by fluorine, chlorine and sulphur occurs

iv. Oxidation also means removal of electropositive element.

e.g.

$2K_4[Fe(CN)_6](aq) + H_2O_2(aq) \rightarrow$

$2K_3[Fe(CN)_6](aq) + 2KOH(aq)$

In this reaction removal of electropositive K occurs

And oxidation state of Fe changes from +2 to +3 that is oxidation takes place

Now overall oxidation is defined as the addition **of oxygen/electronegative element to a substance or removal of hydrogen/ electropositive element from a substance.**

Reduction Reaction:

Reduction reaction can be identify by at least one of the following changes

i. It is the removal of oxygen
ii. Addition of hydrogen.
iii. Removal of electronegative element.
iv. Addition of electropositive element

i. Reduction is the removal of oxygen

e.g.

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

In this example oxidation state of Hg changes from $+2$ to 0 that is reduction takes place

ii. Addition of hydrogen is also a reduction

e.g. $CH_2 = CH_2(g) + H_2(g) \rightarrow H_3C - CH_3(g)$

iii. Reduction is the removal of electronegative element

e.g. $2\,FeCl_3(aq) + H_2(g) \rightarrow 2\,FeCl_2(aq) + 2HCl(aq)$

In this example oxidation state of Fe changes from +3 to +2 that is reduction in oxidation state of Fe

$2HgCl_2(aq) + SnCl_2(aq) \rightarrow Hg_2Cl_2(s) + SnCl_4$

In this reaction simultaneous oxidation of stannous chloride (Sn^{2+}) to stannic chloride (Sn^{4+}) and reduction of Hg^{2+} to Hg^{+} occurs

In the reaction above oxidation and reduction occurs simultaneously hence it is called **redox reaction**

Example:

1. In the reactions given below, identify the species undergoing oxidation and reduction

i) $\mathbf{H_2S(g) + Cl_2(g) \rightarrow 2HCl(g) + S(s)}$

in this reaction oxidation state of S changes from (-2) to (0) that is oxidation of S occurs due to more electronegative element, chlorine is added and more electropositive element, hydrogen is removed and oxidation state of chlorine changes from (0) to (-1) that is reduction occurs

ii) $\mathbf{3Fe_3O_4(s) + 8Al(s) \rightarrow 9Fe(s) + 4Al_2O_3(s)}$
In this reaction aluminium is oxidised because oxygen is added to it O.S. (Oxidation state) of Al changes from (0) to (+3) of and Fe_3O_4 is reduced because oxygen has been removed from it O.S. of Fe changes from $\left(\frac{8}{3}\right)$ to (0)

iii) $\mathbf{2Na\,(s) + H_2(g) \rightarrow 2NaH(s)}$
In this reaction Na is oxidised, O.S. of Na changes from (0) to (+1) and reduction of Hydrogen takes place, O.S. of hydrogen changes from (0) to (-1)

Redox reactions in terms of electron transfer reactions

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
$2Na(s) + S(s) \rightarrow Na_2S(s)$

In each of this reaction above sodium is oxidised due to addition of electronegative elements Chlorine, Oxygen and Sulphur. Simultaneously, chlorine, oxygen and sulphur are reduced because to each of these, the electropositive element sodium has been added.

Sodium chloride, sodium oxide and sodium sulphide are ionic compounds and are written as $Na^+Cl^-(s), (Na^+)_2O^{2-}(s), \text{and } (Na^+)_2S^{2-}(s)$.

In the reaction
$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
O. S. of Na changes from (0) to (+1) i e loss of $2e^-$ from 2Na and O.S. of Cl changes from (0) to (-1) i e gain of $2e^-$ from Cl_2 occurs

In the reaction
$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
O. S. of Na changes from (0) to (+1) i e loss of $4e^-$ from 4Na and O.S. of O changes from (0) to (-2) i e gain of $4e^-$ from O_2 occurs

In the reaction
$2Na(s) + S(s) \rightarrow Na_2S(s)$
O. S. of Na changes from (0) to (+1) i e loss of $2e^-$ from 2Na and O.S. of S changes from (0) to (-2) i e gain of $2e^-$ from S occurs

Each of the above reactions can be written in two separate steps e.g.
$2Na(s) \rightarrow 2Na^+ + 2e^-$ loss of e^- Oxidation Reaction
$Cl_2(g) + 2e^- \rightarrow 2Cl^-$ gain of e^- Reduction Reaction

Each of the above steps known as half reaction which shows the involvement of e^- sum of half reaction gives overall reaction
$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
i.e., half reactions that involve loss of electrons are called **oxidation reactions.** Similarly, the half reactions that involve gain of electrons are called **reduction reactions.**

In the above reactions sodium, which itself oxidised & acts as a **reducing agent** because it donates electron to Cl, O & S and thus helps in reducing them

Whereas Chlorine, oxygen and sulphur are reduced and act as **oxidising agents** because these accept electrons from sodium.

Oxidation: Loss of electron(s) by any species.
Reduction: Gain of electron(s) by any species.
Oxidising agent: Acceptor of electron(s).
Reducing agent: Donor of electron(s).

Competitive Electron Transfer Reactions:

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

In the above two reactions Zn releases electrons to Cu and Cu releases electrons to Ag, so the electron releasing tendency of metals s in the order Zn > Cu > Ag this is also the order of reducing agent.

Oxidation Number

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
In this reaction O.S. of H changes from (0) to (+1) i.e., H_2 get oxidised and O.S. of O changes from (o) to (-2) i.e., O_2 get reduced

$H_2(s) + Cl_2(g) \rightarrow 2HCl(g)$
In this reaction O.S. of H changes from (0) to (+1) i.e., H_2 get oxidised and O.S. of Cl changes from (o) to (-1) i.e., Cl_2 get reduced

$CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(l) + 4HCl(g)$
In this reaction O.S. of C changes from (-4) to (+4) i.e., C get oxidised and O.S. of Cl changes from (o) to (-1) i.e., Cl_2 get reduced

Oxidation number denotes the oxidation state of an element in a compound ascertained according to a set of rules formulated on the basis that electron pair in a covalent bond belongs entirely to more electronegative element.

Rules for calculation of Oxidation number

Rule-1

In elements, in the free or the uncombined state, each atom bears an oxidation number of zero. Evidently each atom in $H_2, O_2, Cl_2, O_3, P_4, S_8, Na, Mg, Al$ has the oxidation number zero

Rule-2

For ions composed of only one atom, the oxidation number is equal to the charge on the ion.

Thus Na^+ ion has an oxidation number of +1, Mg^{2+} ion, +2, Fe^{3+} ion, +3, Cl^- ion, –1, O^{2-} ion, –2; and so on.

In their compounds all alkali metals have oxidation number of +1, and all alkaline earth metals have an oxidation number of +2. Aluminium is regarded to have an oxidation number of +3 in all its compounds.

Rule-3

The oxidation number of oxygen in most compounds is –2. However, in the case of peroxides and superoxide, the compounds of oxygen in which oxygen atoms are directly linked to each other.

While in peroxides (e.g., H_2O_2, Na_2O_2), each oxygen atom is assigned an oxidation

Number of –1,

In superoxide's (e.g., KO_2, RbO_2) each oxygen atom is assigned an oxidation number of – (½).

When oxygen is bonded to fluorine in such compounds e.g., oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), the oxygen is assigned an oxidation number of +2 and +1, respectively.

The number assigned to oxygen will depend upon the bonding state of oxygen.

Rule – 4

The oxidation number of hydrogen is +1, except when it is bonded to metals in binary compounds (that is compounds containing two elements). For example, in LiH, NaH, and CaH_2, its oxidation number is –1.

Rule – 5

In all its compounds, fluorine has an oxidation number of –1. Other halogens (Cl, Br, and I) also have an oxidation number of –1, when they occur as halide ions in their compounds.

Chlorine, bromine and iodine when combined with oxygen, for example in oxoacids and oxoanions, have positive oxidation numbers.

Rule – 6

The algebraic sum of the oxidation number of all the atoms in a compound must be zero. In polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must equal the charge on the ion.

e.g. In Thus, $(CO_3)^{2-}$ the sum of oxidation number of three oxygen atoms and one carbon atom, must equal to –2.

Important Generalisations

1. Metallic elements have positive oxidation number and non-metallic elements have positive or negative oxidation number.
2. The atoms of transition elements usually display several positive oxidation states
3. The highest oxidation number of a representative element (S & P Block) is the group number for the first two groups (for S block) and the group number minus 10 (for P Block) for the other groups.
4. Thus the highest value of oxidation number exhibited by an atom of an element generally increases across the period in the periodic table.
   In the third period, the highest value of oxidation number changes from 1 to 7 as indicated in the table

Group	**1**	**2**	**13**	**14**	**15**	**16**	**17**
Element	Na	Mg	Al	Si	P	S	Cl
Compound	NaCl	$MgSO_4$	AlF_3	$SiCl_4$	P_4O_{10}	SF_6	$HClO_4$
O.S.	+1	+2	+3	+4	+5	+6	+7

Summary:

Oxidation: An increase in the oxidation number of the element in the given substance.

Reduction: A decrease in the oxidation number of the element in the given substance.

Oxidising agent (Oxidant): A reagent which can increase the oxidation number of an element in a given substance. These reagents are called as oxidants also.

Reducing agent (Reductant): A reagent which lowers the oxidation number of an element in a given substance. These reagents are also called as reductants.

Redox reactions: Reactions which involve change in oxidation number of the interacting species.

Types of Redox Reactions

1) Combination reactions
2) Decomposition reactions

3) Displacement reactions
   a) Metal displacement
   b) Non-metal displacement
4) Disproportionation reactions

1) Combination Reactions

$A + B \rightarrow C$

Either A and B or both A and B must be in the elemental form for such a reaction to be a redox reaction

e.g. $C(s) + O_2(g) \rightarrow CO_2(g)$

In this reaction O.S. of C changes from (0) to (+4) i.e. C get Oxidised and acts as reducing agent (Reductant), O.S. of Oxygen changes from (0) to (-2) i.e. Oxygen get reduced and acts as Oxidising agent (Oxidant)

$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

In this reaction O.S. of Mg changes from (0) to (+2) i.e. Mg get Oxidised and acts as reducing agent (Reductant), O.S. of Nitrogen changes from (0) to (-3) i.e. nitrogen get reduced and acts as Oxidising agent (Oxidant)

$CH_4(g) + 2O_2(g) \rightarrow CO2(g) + 2H_2O(l)$

In this reaction O.S. of C changes from (-4) to (+4) i.e. C get Oxidised and acts as reducing agent (Reductant), O.S. of Oxygen changes from (0) to (-2) i.e. oxygen get reduced and acts as Oxidising agent (Oxidant)

2) Decomposition Reactions

Decomposition reaction leads to the breakdown of a compound into two or more components at least one of which must be in the elemental state.

$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

In this reaction O.S. of O changes from (-2) to (0) and O.S. of H changes from (+1) to (0)

$2NaH(s) \rightarrow 2Na(s) + H_2(g)$

In this reaction O.S. of Na changes from (+1) to (0) and O.S. of H changes from (-1) to (0)

$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

In this reaction O.S. of Cl changes from (+5) to (-1) and O.S. of Oxygen changes from (-6) to (0)

Noted that all decomposition reactions are not redox reactions. For example, decomposition of Calcium carbonate is not a redox reaction.

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

In this reaction O.S. of Ca, C & O remains (+2), (+4) & (-2) respectively

3) Displacement Reactions

$X + YZ \rightarrow XZ + Y$

Displacement reactions have two type's namely metal displacement and non-metal displacement.

a. Metal Displacement

A metal in a compound can be displaced by another metal in the uncombined state

e.g.

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

In this reaction O.S. of Cu changes from (+2) to (0) and O.S. of Zn changes from (0) to (+2)

$V_2O_5(s) + 5Ca(s) \rightarrow 2V(s) + 5CaO(s)$

In this reaction O.S. of V changes from (+5) to (0) and O.S. of Ca changes from (0) to (+2)

$TiCl_4(l) + 2Mg(s) \rightarrow Ti(s) + 2MgCl_2(s)$

In this reaction O.S. of Ti changes from (+4) to (0) and O.S. of Mg changes from (0) to (+2)

$Cr_2O_3(s) + 2Al(s) \rightarrow Al_2O_3(s) + 2Cr(s)$

b. Non-metal Displacement

The non-metal displacement redox reactions include hydrogen displacement and a rarely occurring reaction involving oxygen displacement. All alkali metals and some alkaline earth metals (Ca, Sr, and Ba) which are very good reductants, will displace hydrogen from cold water.

e.g.

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

In this reaction O.S. of Na changes from (0) to (+1) and O.S. of H changes from (+1) to (0)

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

In this reaction O.S. of Ca changes from (0) to (+2) and O.S. of H changes from (+1) to (0)

Less active metals such as magnesium and iron react with steam to produce dihydrogen gas

$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g)$

In this reaction O.S. of Mg changes from (0) to (+2) and O.S. of H changes from (+1) to (0)

$2Fe(s) + 3H_2O(l) \rightarrow Fe_2O_3(s) + 3H_2(g)$

In this reaction O.S. of Fe changes from (0) to (+3) and O.S. of H changes from (+1) to (0)

Many metals, including those which do not react with cold water, are capable of displacing hydrogen from acids.

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
In this reaction O.S. of Zn changes from (0) to (+2) and O.S. of H changes from (+1) to (0)

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
In this reaction O.S. of Mg changes from (0) to (+2) and O.S. of H changes from (+1) to (0)

$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
In this reaction O.S. of Fe changes from (0) to (+2) and O.S. of H changes from (+1) to (0)
From the above reactions we can predict the reactivity of metals towards the $H_2(g)$ evolution

$2H_2O\,(l) + 2F_2(g) \rightarrow 4HF(aq) + O_2(g)$
In this reaction O.S. of O changes from (-2) to (0) and O.S. of F changes from (0) to (-1)

$Cl_2(g) + 2KBr(aq) \rightarrow 2KCl\,(aq) + Br_2(l)$
In this reaction O.S. of Cl changes from (0) to (-1) and O.S. of Br changes from (-1) to (0)

$Cl_2(g) + 2KI(aq) \rightarrow 2KCl\,(aq) + I_2(s)$
In this reaction O.S. of Cl changes from (0) to (-1) and O.S. of I changes from (-1) to (0)

$Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(l)$
In this reaction O.S. of Cl changes from (0) to (-1) and O.S. of Br changes from (-1) to (0)

$Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(s)$
In this reaction O.S. of Cl changes from (0) to (-1) and O.S. of I changes from (-1) to (0)

$Br_2(l) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(s)$
In this reaction O.S. of Br changes from (0) to (-1) and O.S. of I changes from (-1) to (0)

4) Disproportionation Reactions

In a disproportionation reaction an element in one oxidation state is simultaneously oxidised and reduced. One of the reacting substances in a disproportionation reaction always contains an element that can exist in at least three oxidation states. The element in the form of reacting substance is in the intermediate oxidation state; and both higher and lower oxidation states of that element are formed in the reaction.
e.g.
$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$
Here the oxygen of peroxide, which is present in –1 state, is converted to zero oxidation state in O_2 and decreases to –2 oxidation state in H_2O.
Phosphorous, sulphur and chlorine undergo disproportionation in the alkaline medium as shown

$P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-$
Here the P, which is present in (0) state, is converted to (-3) oxidation state in PH_3 and (-1) oxidation state in $H_2PO_2^-$.

$S_8(s) + 12OH^-(aq) \rightarrow 4S^{2-}(aq) + 2S_2O_3^{2-}(aq) + 6H_2O(l)$
Here the S, which is present in (0) state, is converted to (-2) oxidation state in S^{2-} and (+2) oxidation state in $S_2O_3^{2-}$.

$Cl_2(g) + 2OH^-(aq) \rightarrow ClO^-(aq) + Cl^-(aq) + H_2O(l)$
Here the Cl , which is present in (0) state, is converted to (+1) oxidation state in ClO^- and (-1) oxidation state in Cl^-.

$2F_2(g) + 2OH^-(aq) \rightarrow 2F^-(aq) + OF_2(g) + H_2O(l)$
Here the F, which is present in (0) state, is converted to (-1) oxidation state in F^- and (-1) oxidation state in OF_2. This departure shown by fluorine is not surprising for us as we know the limitation of fluorine that, being the most electronegative element, it cannot exhibit any positive oxidation state. This means that among halogens, fluorine does not show a disproportionation tendency.

Balancing of Redox Reactions

1) Oxidation Number Method:
2) Half Reaction Method:

1) Oxidation Number Method:

Step 1: Write the correct formula for each reactant and product.

Step 2: Identify atoms which undergo change in oxidation number in the reaction by assigning the oxidation number to all elements in the reaction.

Step 3: Calculate the increase or decrease in the oxidation number per atom and for the entire molecule/ion in which it occurs. If these are not equal then multiply by suitable number so that these become equal.

Step 4: Ascertain the involvement of ions if the reaction is taking place in water, adds H^+ or OH^- ions to the expression on the appropriate side so that the total ionic charges of reactants and products are equal. If the reaction is carried out in acidic solution, use H^+ ions in the equation; if in basic solution, use OH^- ions.

Step 5: Make the numbers of hydrogen atoms in the expression on the two sides equal by adding water (H_2O) molecules to the reactants or products. Now, also check the number of oxygen atoms. If there are the same number of oxygen atoms in the reactants and products, the equation then represents the balanced redox reaction.

e.g. Permanganate ion reacts with bromide ion in basic medium to give manganese dioxide and bromate ion. Write the balanced ionic equation for the reaction.

Step 1:

$$MnO_4^-(aq) + Br^-(aq) \rightarrow MnO_2(s) + BrO_3^-(aq)$$

Step 2: Assign oxidation numbers for Mn and Br

$$\underset{+7}{MnO_4^-(aq)} + \underset{-1}{Br^-(aq)} \rightarrow \underset{+4}{MnO_2(s)} + \underset{+5}{BrO_3^-(aq)}$$

this indicates that permanganate ion is the oxidant and bromide ion is the reductant.

Step 3: Calculate the increase and decrease of oxidation number, and make the increase equal to the decrease

$$\underset{+7}{2MnO_4^-(aq)} + \underset{-1}{Br^-(aq)} \rightarrow \underset{+4}{2MnO_2(s)} + \underset{+5}{BrO_3^-(aq)}$$

Step 4: As the reaction occurs in the basic medium, and the ionic charges are not equal on both sides, add $2OH^-$ ions on the right to make ionic charges equal

$$\underset{+7}{2MnO_4^-(aq)} + \underset{-1}{Br^-(aq)} \rightarrow \underset{+4}{2MnO_2(s)} + \underset{+5}{BrO_3^-(aq)} + 2OH^-$$

Step 5: Finally, count the hydrogen atoms and add appropriate number of water molecules (i.e. one H_2O molecule) on the left side to achieve balanced redox change.

$$2MnO_4^-(aq) + Br^-(aq) + H_2O\ (l) \rightarrow 2MnO_2(s) + BrO_3^-(aq) + 2OH^-$$

e.g. Write the net ionic equation for the reaction of potassium dichromate(VI), $K_2Cr_2O_7$ with sodium sulphite, Na_2SO_3, in an acid solution to give chromium(III) ion and the sulphate ion.

Step 1: The skeletal ionic equation is

$$Cr_2O_7^{2-}(aq) + SO_3^{2-}(aq) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$$

Step 2: Assign oxidation numbers for Cr and S

$$\underset{+6}{Cr_2O_7^{2-}(aq)} + \underset{+4}{SO_3^{2-}(aq)} \rightarrow \underset{+3}{Cr^{3+}(aq)} + \underset{+6}{SO_4^{2-}(aq)}$$

Step 3: Calculate the increase and decrease of oxidation number, and make them equal: from step-2 we can notice that there is change in oxidation state of chromium and sulphur. Oxidation state of chromium changes form +6 to +3. There is decrease of +3 in oxidation state of chromium on right hand side of the equation. Oxidation state of sulphur changes from +4 to +6. There is an increase of +2 in the oxidation state of sulphur on right hand side. To make the increase and decrease of oxidation state equal, place numeral 2 before chromium ion on right hand side and numeral 3 before sulphate ion on right hand side and balance the chromium and sulphur atoms on both the sides of the equation.

$$\underset{+6}{Cr_2O_7^{2-}(aq)} + \underset{+4}{3SO_3^{2-}(aq)} \rightarrow \underset{+3}{2Cr^{3+}(aq)} + \underset{+6}{3SO_4^{2-}(aq)}$$

Step 4: As the reaction occurs in the acidic medium, and further the ionic charges are not equal on both the sides, add $8H^+$ on the left to make ionic charges equal

$$Cr_2O_7^{2-}(aq) + 3SO_3^{2-}(aq) + 8H^+ \rightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq)$$

Step 5: Finally, count the hydrogen atoms, and add appropriate number of water molecules (i.e., $4H_2O$) on the right to achieve balanced redox change.

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