

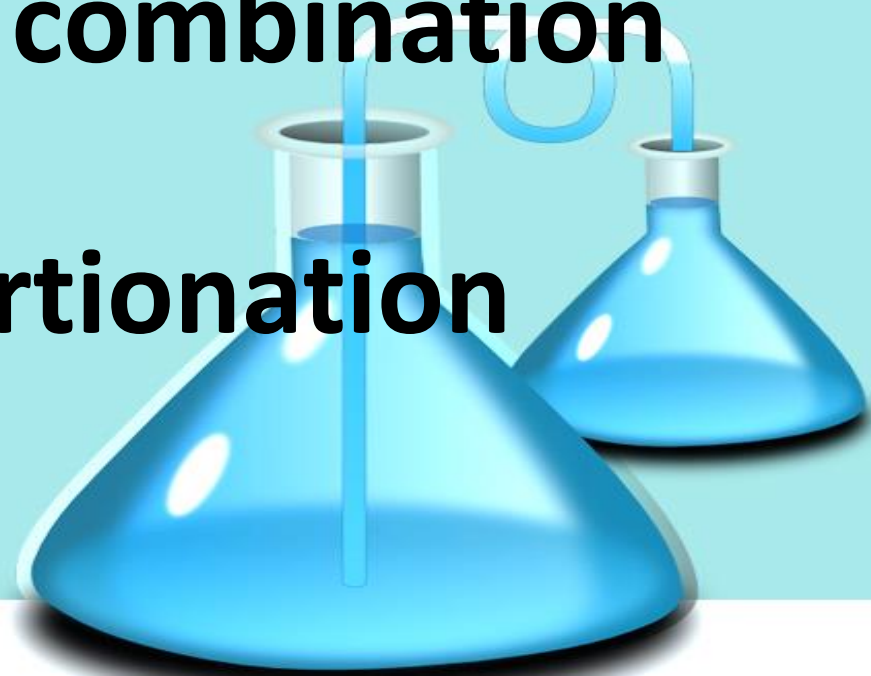
**CLASS XI
CHEMISTRY**

REDOX REACTIONS

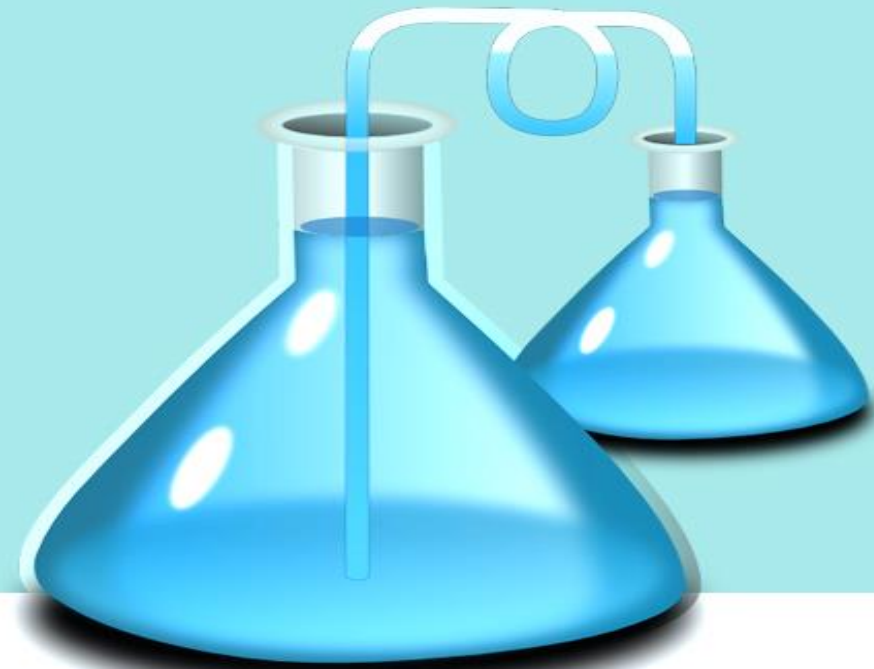


OBJECTIVES

- Use the concept of oxidation number to identify oxidant and reductant in a reaction
- Classify redox reaction into combination (synthesis), decomposition, displacement and disproportionation reactions;



OXIDATION NUMBER



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.



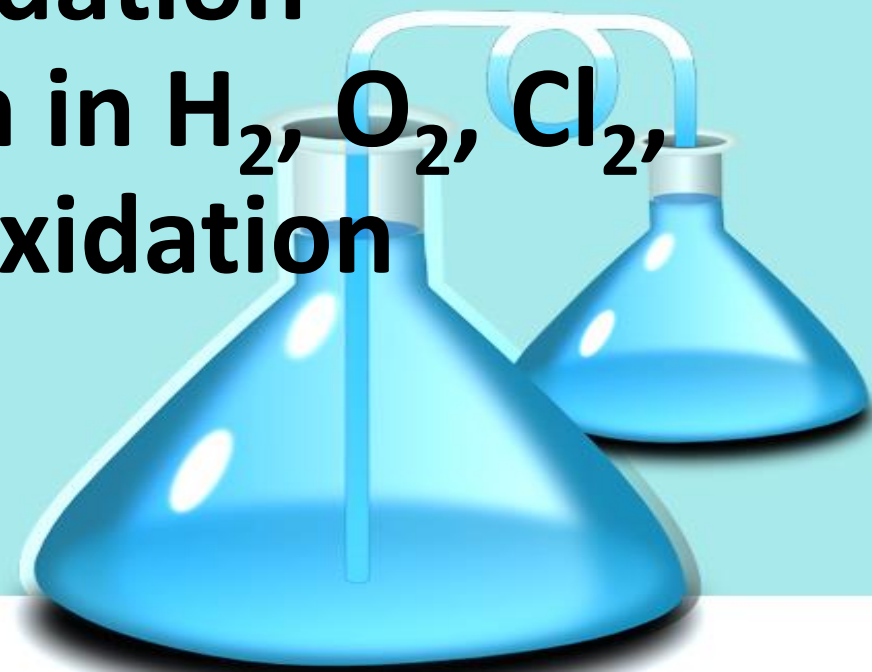
In this method, an important assumption is that there is a complete transfer of electron or electrons from a less electronegative atom to a more electronegative atom.

If two or more than two atoms of an element are present in the molecule or ion such as $\text{K}_2\text{Cr}_2\text{O}_7$ or SO_4^{2-} , then oxidation number of the atom of that element will then be the average of the oxidation number of all the atoms of that element.



Rules for calculation of oxidation number

(1) In elements, in the free or the uncombined state, or in any allotropic forms, each atom bears an oxidation number of zero. So, each atom in H_2 , O_2 , Cl_2 , O_3 , P_4 , S_8 , Na, Mg, Al has the oxidation number zero.



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

Oxidation number of K^+ is +1, Ca^{2+} ion is +2, Fe^{2+} ion is +3

Oxidation number of Cl^- ion is -1, O^{2-} ion is -2



All alkali metals have oxidation number of +1 in their compounds.

All alkaline earth metals have an oxidation number of +2 in their compounds.

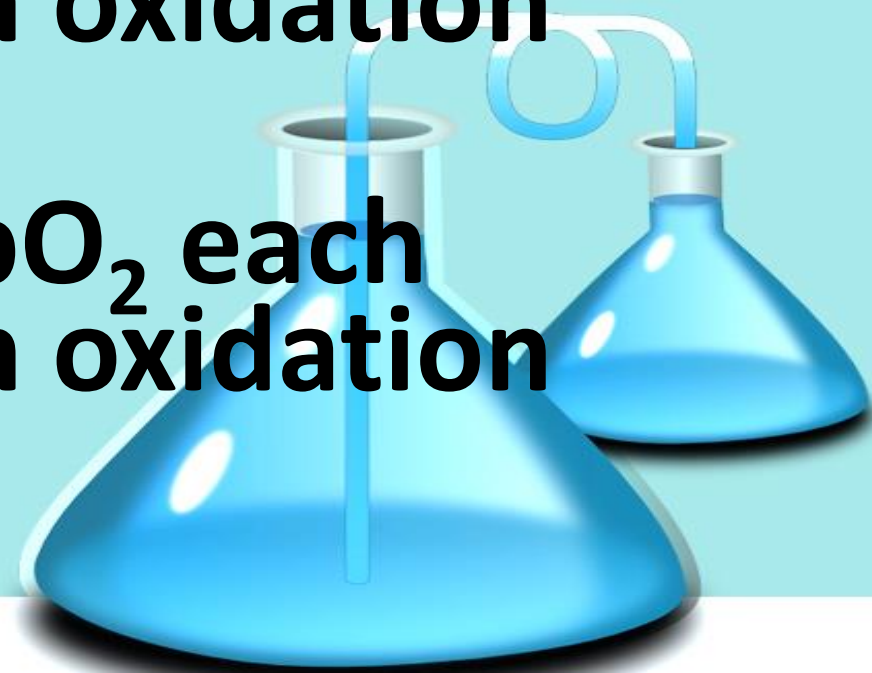
Aluminium has an oxidation number of +3 in all its compounds.



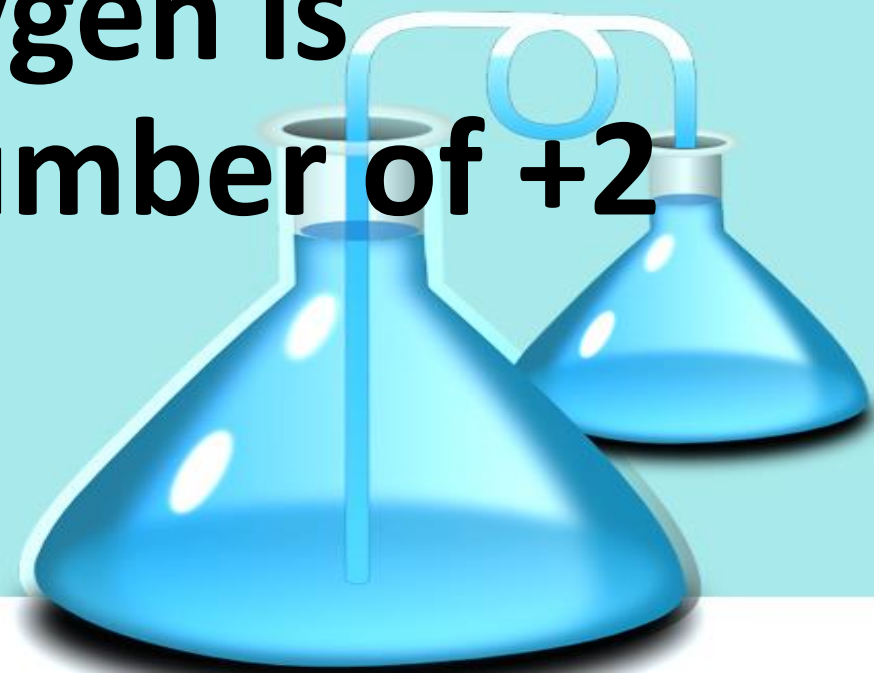
(3) The oxidation number of oxygen in most compounds is -2 .

Exceptions: in case of peroxides and superoxides.

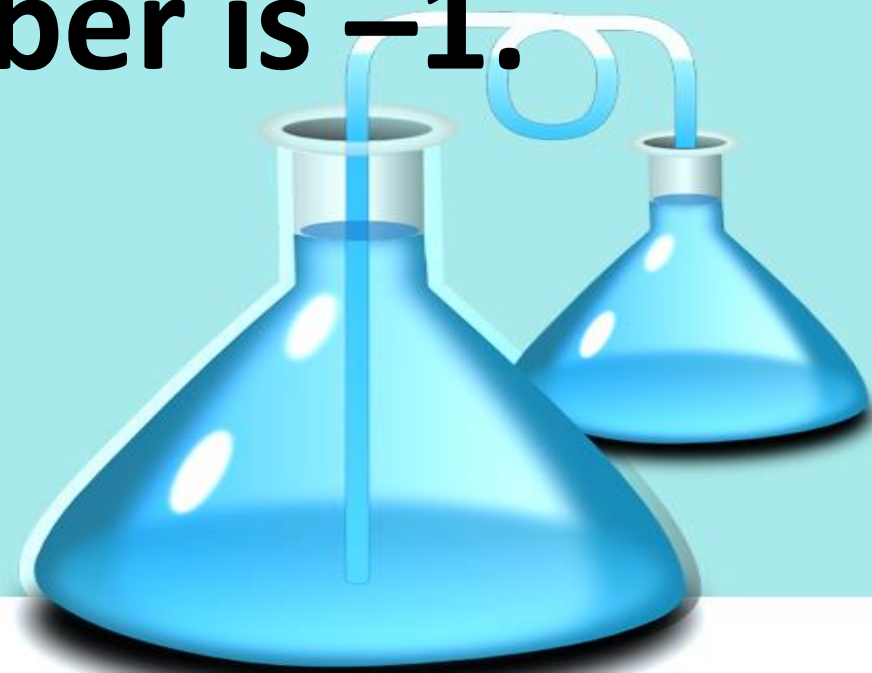
- In peroxides, like H_2O_2 , Na_2O_2 , each oxygen atom is assigned an oxidation number of -1
- In superoxides like, KO_2 , RbO_2 each oxygen atom is assigned an oxidation number of $-(\frac{1}{2})$



Another exception occurs 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.



(4) The oxidation number of hydrogen is +1, except when it is bonded to metals in binary compounds. E.g., in LiH, NaH, and CaH_2 , the oxidation number is -1.



(5) In all its compounds, fluorine, the most electronegative atom, always 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 in oxoacids and oxoanions, have positive oxidation numbers.



(6) In binary compounds of metal and non-metal, the oxidation number of the metal is always positive and the oxidation number of non-metal is always negative.



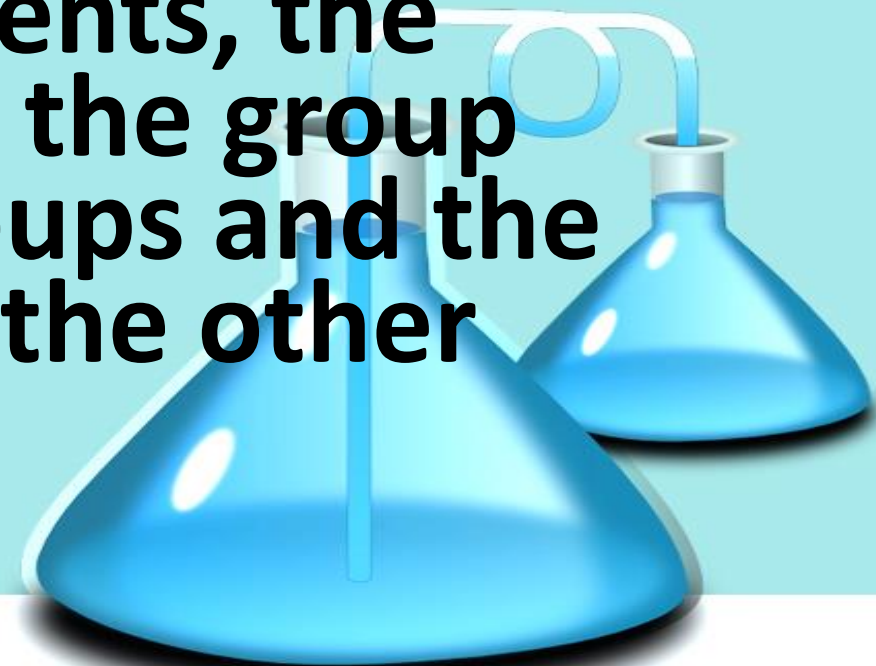
(7) The algebraic sum of the oxidation number of all the atoms in a compound is zero.



(8) In polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion is equal to the charge on the ion. E.g., the sum of oxidation number of four oxygen atoms and one sulphur atom in the sulphate ion, $(\text{SO}_4)^{2-} = -2$.



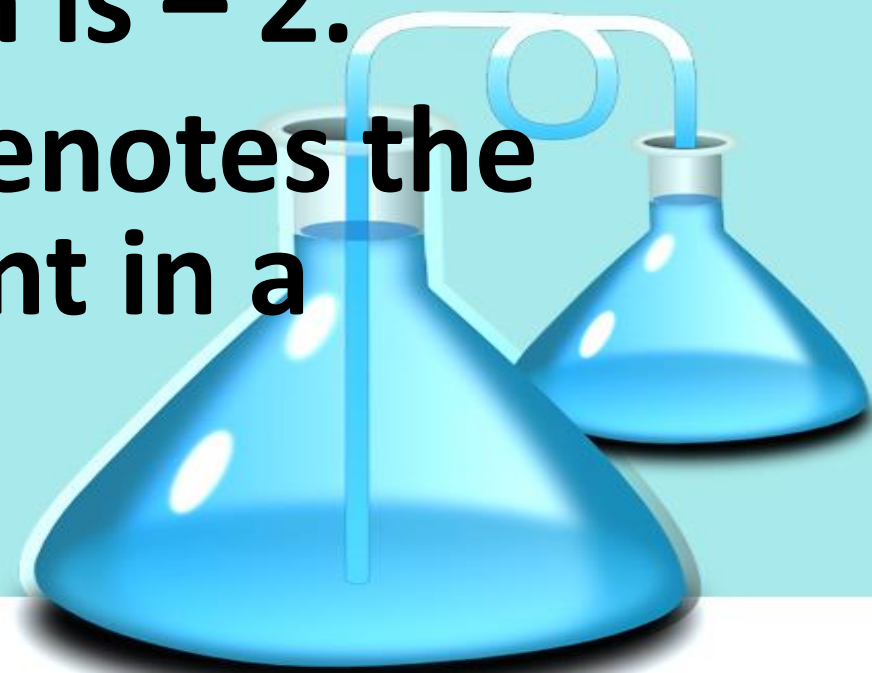
- **Metallic elements have positive oxidation number and non-metallic elements have positive or negative oxidation number.**
- **The atoms of transition elements usually display several positive oxidation states.**
- **For the representative elements, the highest oxidation number is the group number for the first two groups and the group number minus 10 for the other groups.**



Consider CO₂. In this compound, the oxidation state of carbon is +4, that is also its oxidation number.

Likewise, the oxidation state as well as oxidation number of oxygen is – 2.

So, the oxidation number denotes the oxidation state of an element in a compound.



Now let us apply the above rules to calculate the oxidation number of element in the molecule or ion.

Calculate the oxidation number of Sulphur in Sodium thiosulphate
 $\text{Na}_2\text{S}_2\text{O}_3$



$\text{Na}_2\text{S}_2\text{O}_3$ Sodium thiosulfate

+1

x

-2

Na_2

S_2

O_3

$(+1 \times 2)$

$(x \times 2)$

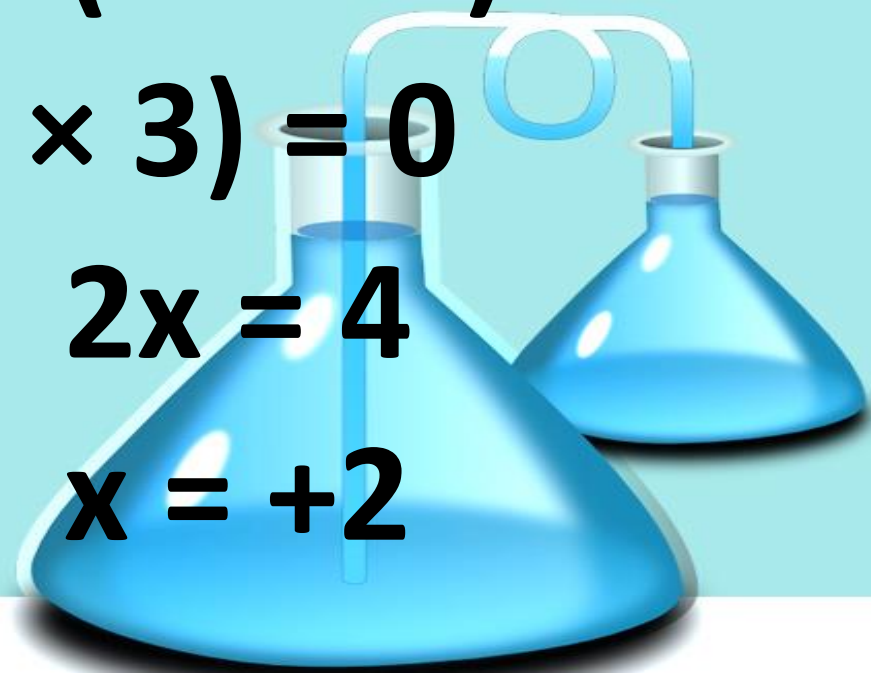
(-2×3)

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

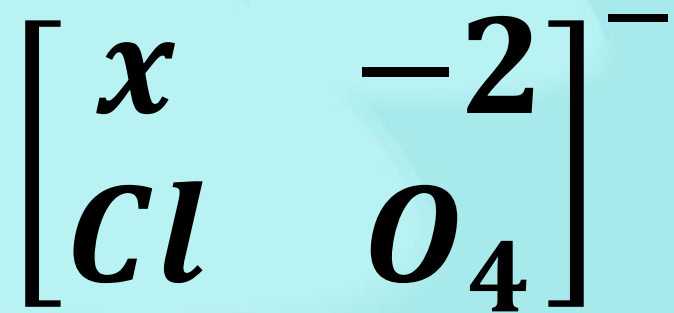
$+2 + 2x - 6 = 0$

or $2x = 4$

$x = +2$

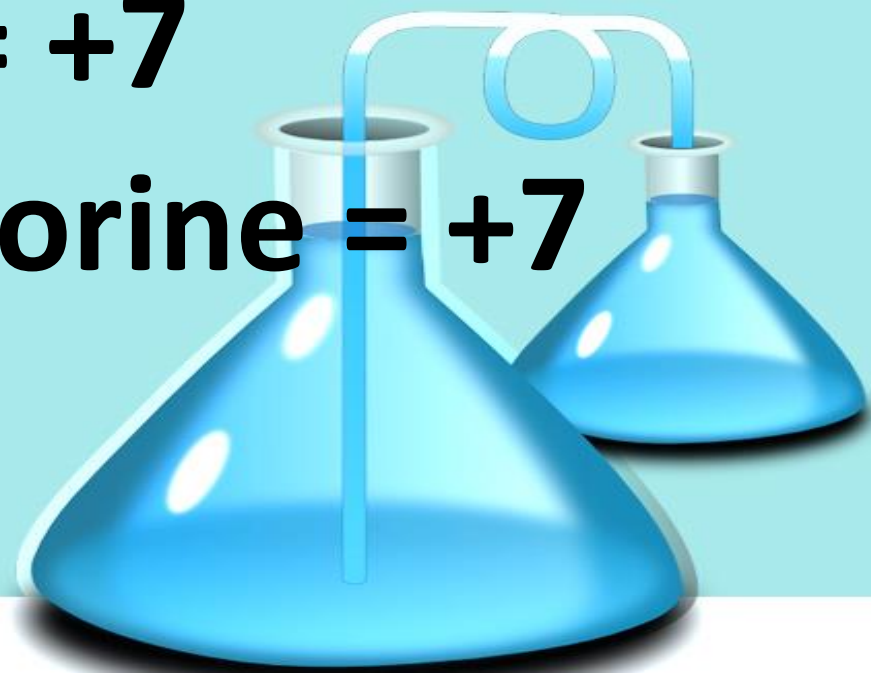


Calculate the oxidation number of Chlorine in (Perchlorate Ion), $(\text{ClO}_4)^-$



$$x + (-2 \times 4) = -1 \text{ or } x = +7$$

Oxidation number of Chlorine = +7



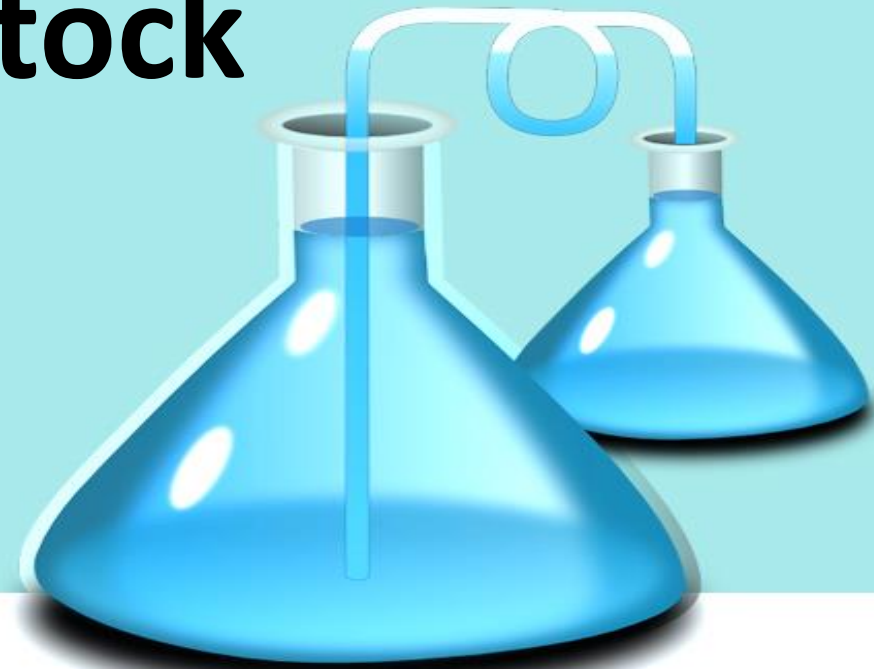
**Now, similarly calculate the
oxidation number of**

- (i) Sulphur in Sulphuric Acid**
- (ii) Manganese in Potassium Permanganate**

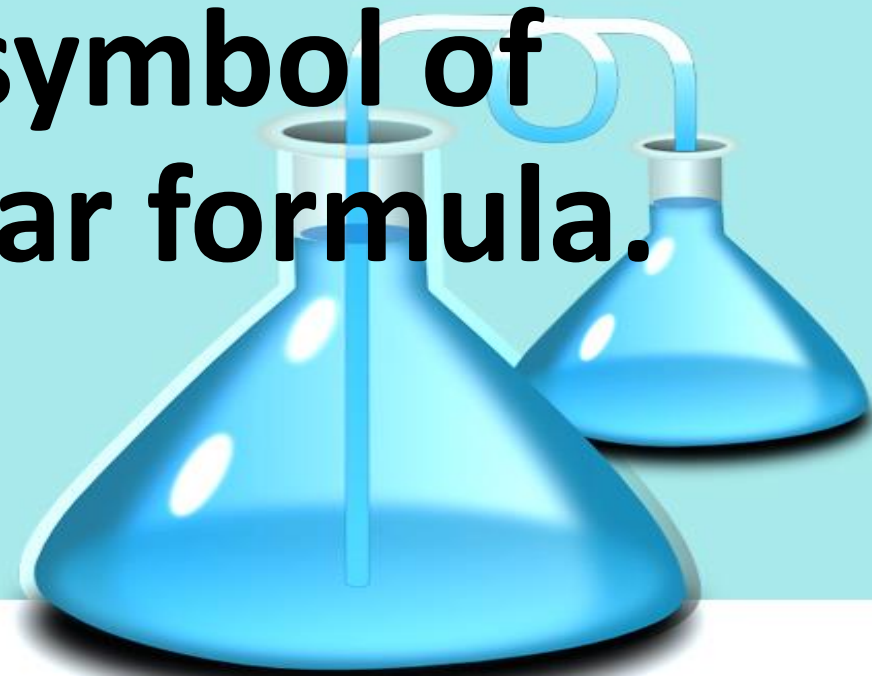


STOCK NOTATION

This notation was given by German chemist, Alfred Stock



According to Stock notation, the oxidation number is expressed by putting a Roman numeral representing the oxidation number in parenthesis after the symbol of the metal in the molecular formula.



STOCK NOTATION

Stannous chloride is written as
 Sn(II)Cl_2

Stannic chloride is written as
 Sn(IV)Cl_4

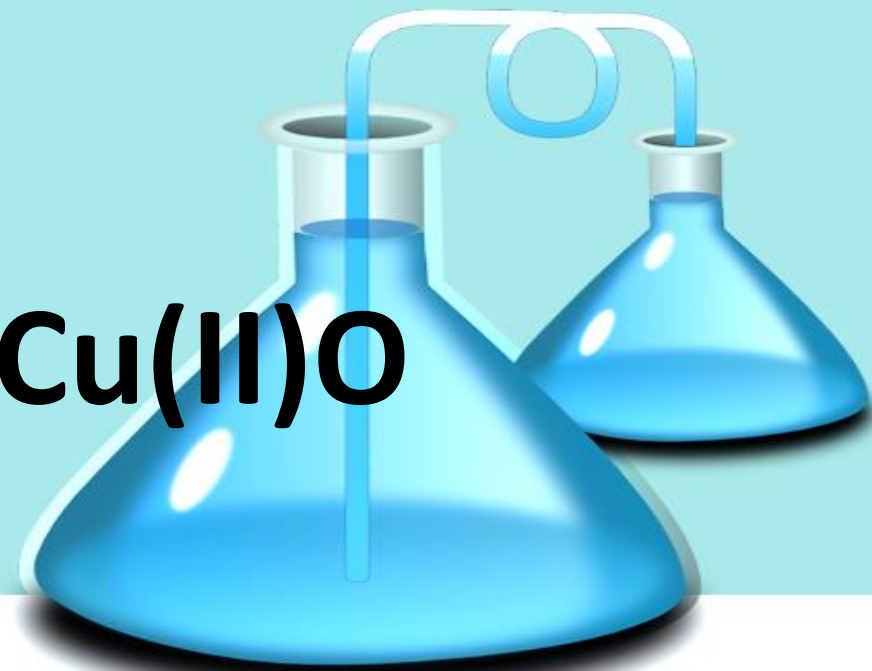


Using Stock notation, represent the following compounds

FeO, Fe₂O₃, CuI, CuO

The Stock notation is

Fe(II)O, Fe₂(III)O₃, Cu(I)I, Cu(II)O



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 or Oxidant:

A reagent which can increase the oxidation number of an element in a given substance.

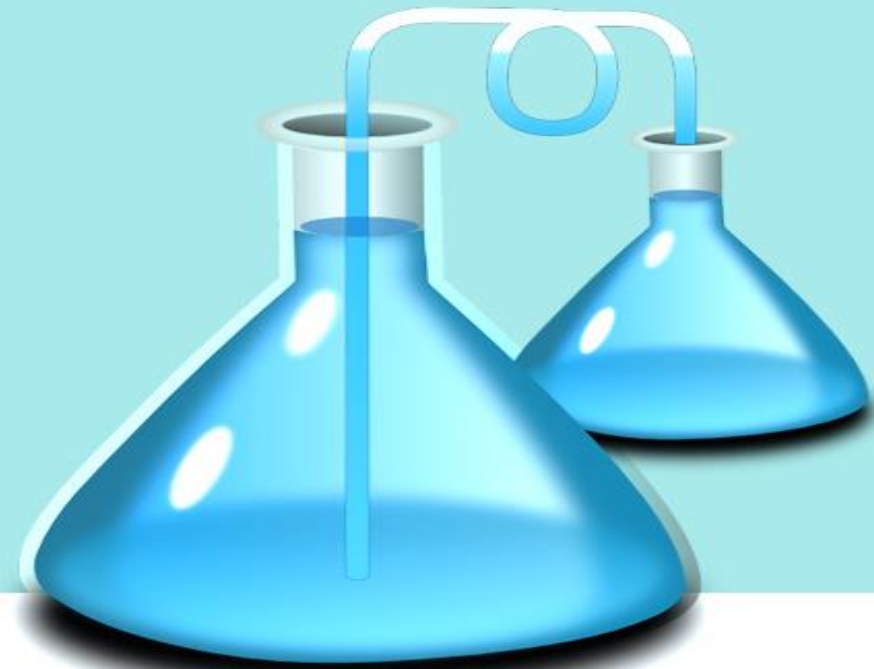
Reducing agent or Reductant:

A reagent which can decrease the oxidation number of an element in a given substance.

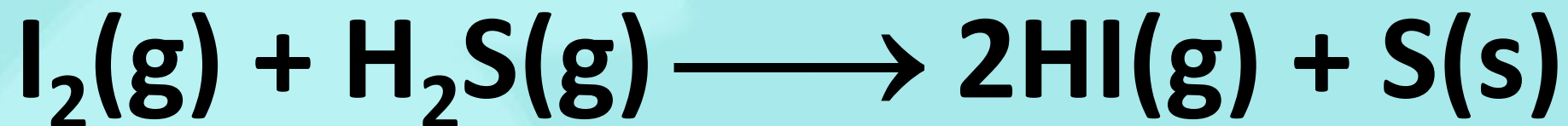


Redox reactions

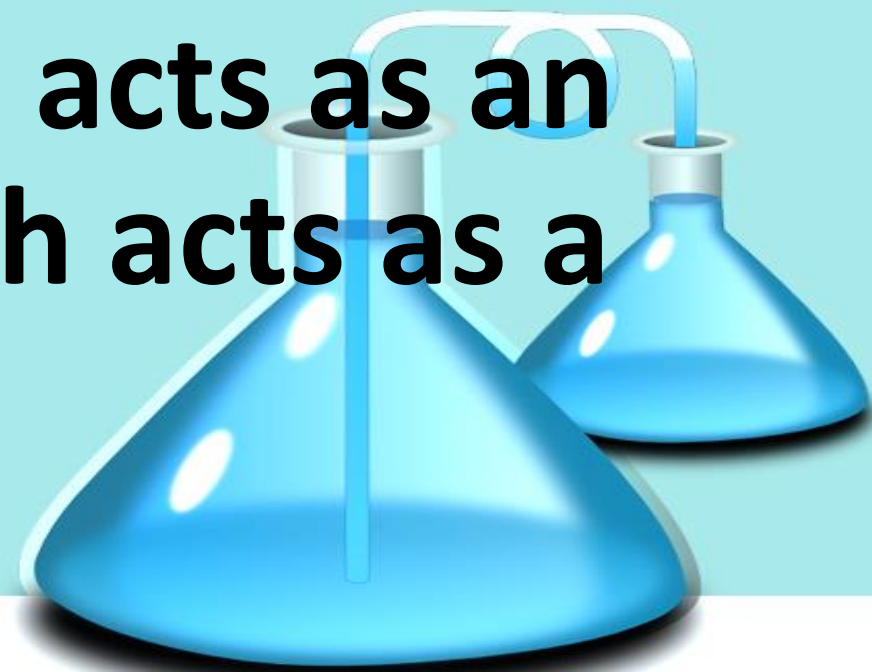
Reactions which involve change in oxidation number of the interacting species.

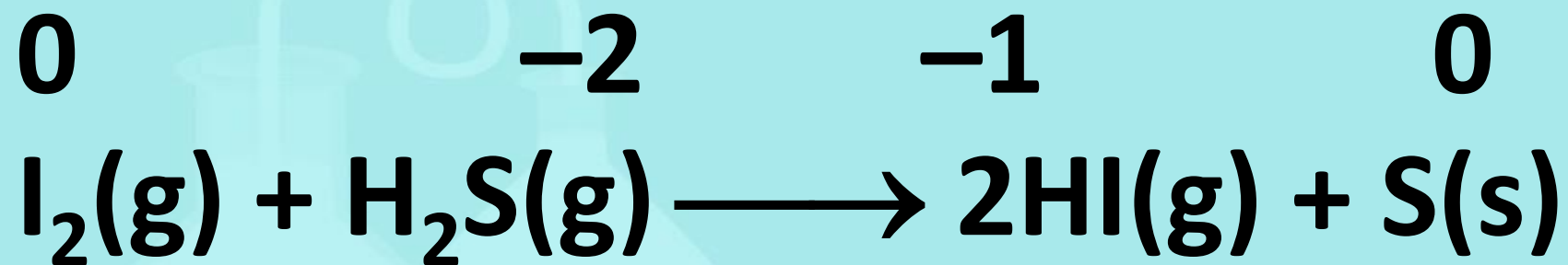


Is the following reaction is a redox reaction?

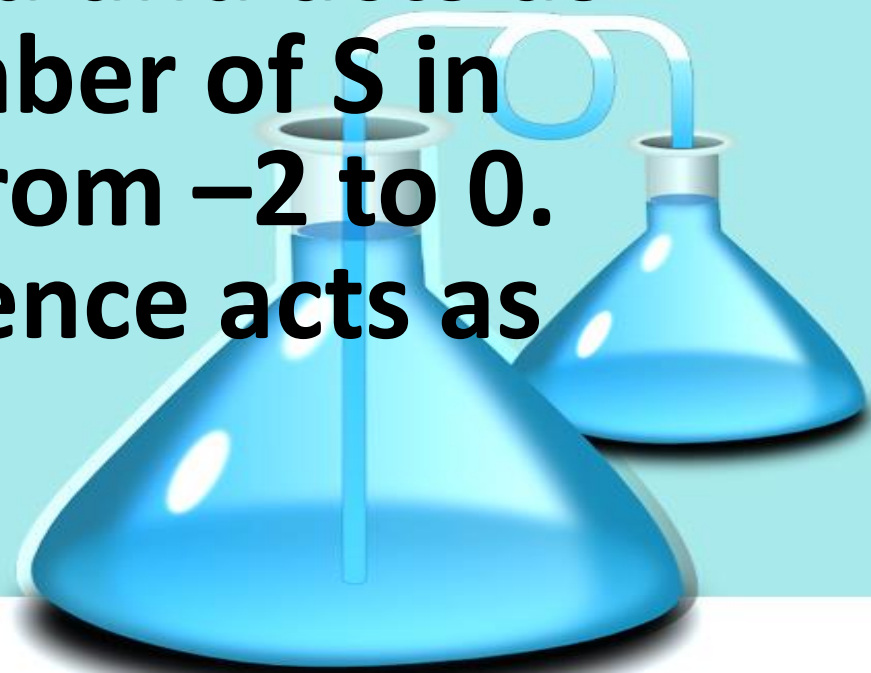


Also, identify the species oxidised/reduced, which acts as an oxidising agent and which acts as a reducing agent.





Oxidation number of iodine decreases from 0 to -1. Therefore, it is reduced and acts as an oxidant. The oxidation number of S in hydrogen sulphide increases from -2 to 0. Therefore, it is oxidized and hence acts as reductant.



Types of Redox Reactions

- **Combination reactions**
- **Decomposition reactions**
- **Displacement reactions**
- **Disproportionation reactions**



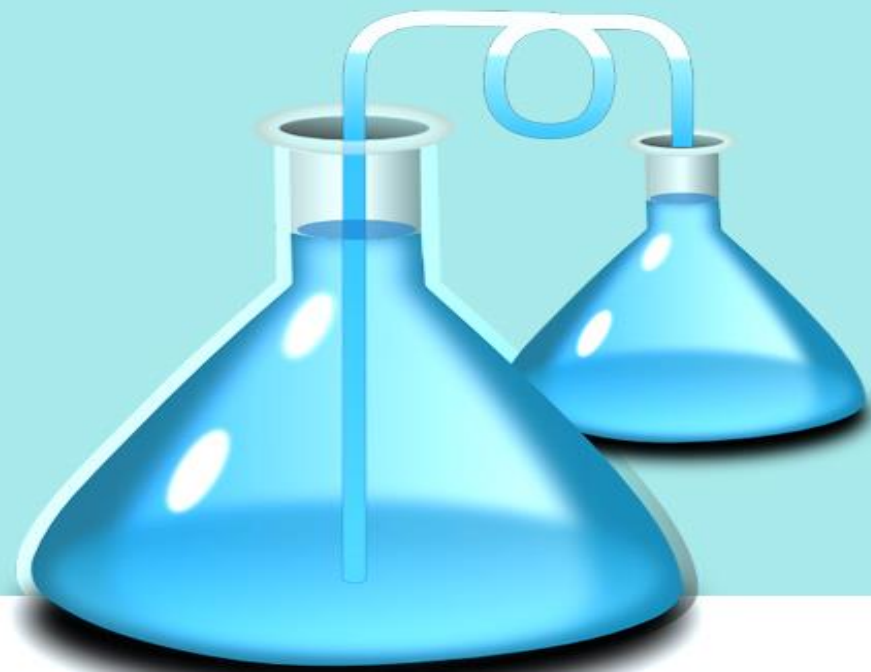
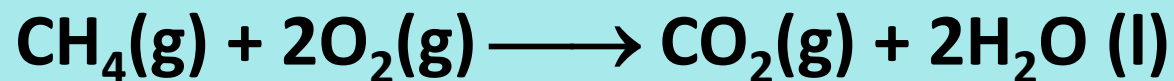
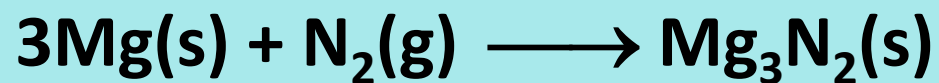
Combination reactions

A combination reaction may be represented as



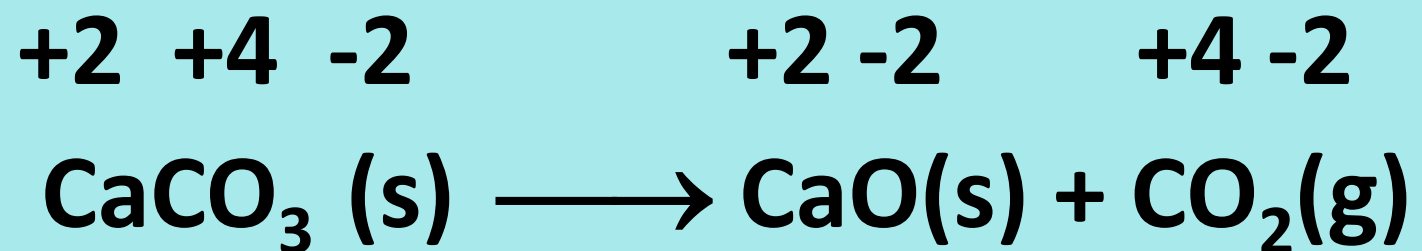
For combination reaction to be a redox reaction, either A and B or both A and B must be in the elemental form.

All combustion reactions are redox reactions.



Decomposition reactions

For a decomposition reaction to be a redox reaction, a decomposition reaction should lead to the breakdown of a compound into two or more components, at least one of the component must be in the elemental form.



Displacement reactions

In a displacement reaction, an ion or an atom in a compound is replaced by an ion or an atom of another element. It may be represented as:

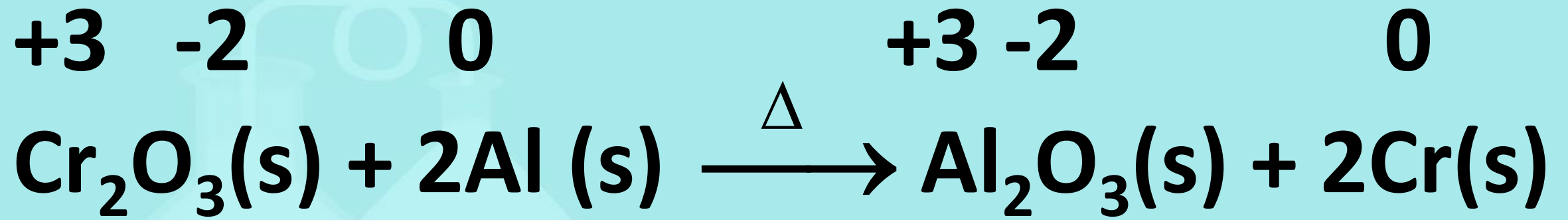


Classification of Displacement reactions:

(i) metal displacement

(ii) non-metal displacement

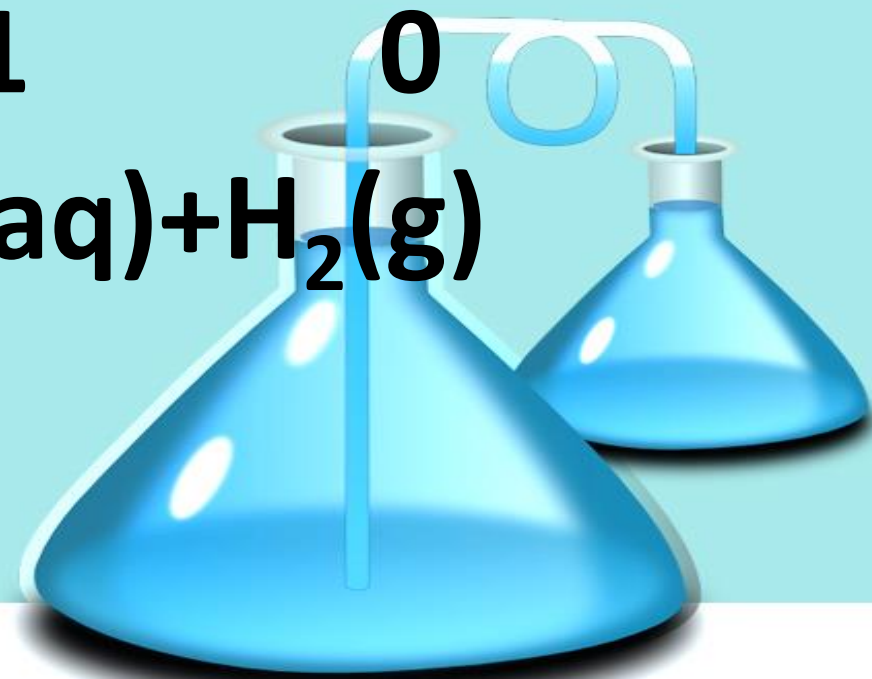
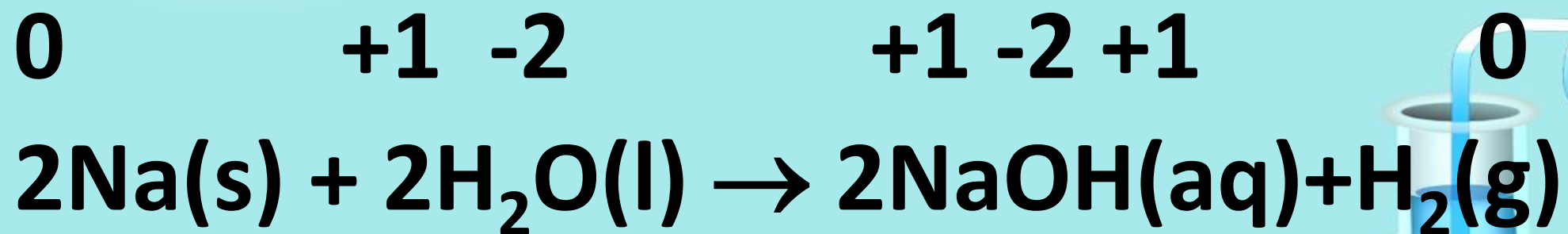




Application of metal displacement reactions is seen in metallurgical processes in which pure metals are obtained from their compounds in ores.



(b) Non-metal displacement: The non-metal displacement redox reactions include hydrogen displacement



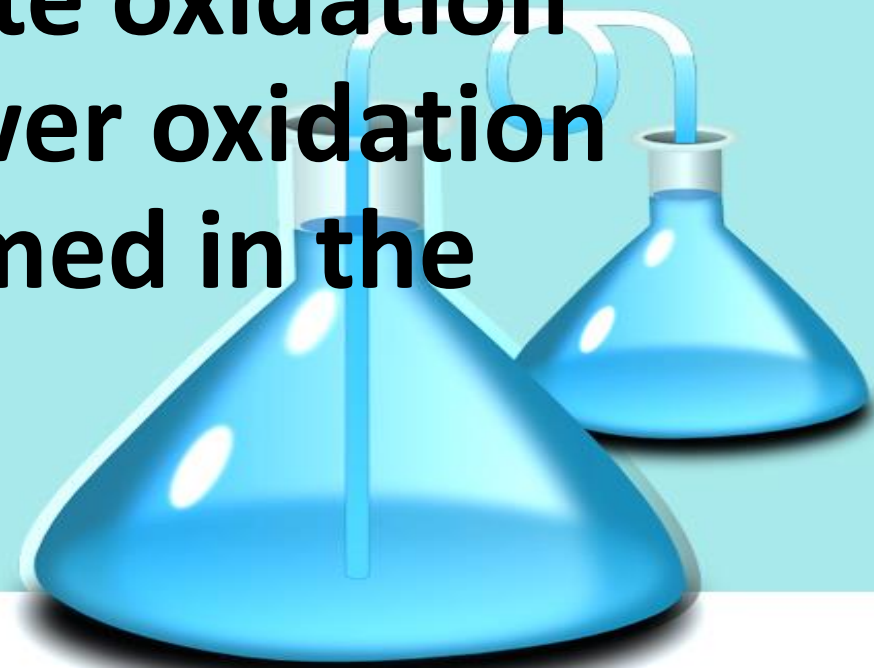
Disproportionation reactions

In a disproportionation reaction, an element in one oxidation state is simultaneously oxidised and reduced.

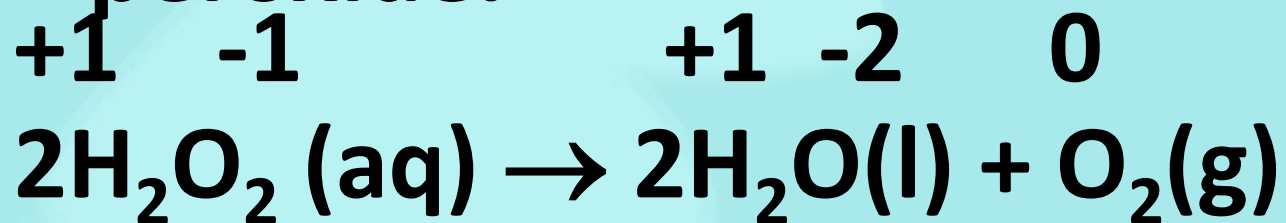


So, one of the reacting substances in a disproportionation reaction always contains an element that can exist in at least three oxidation states.

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.

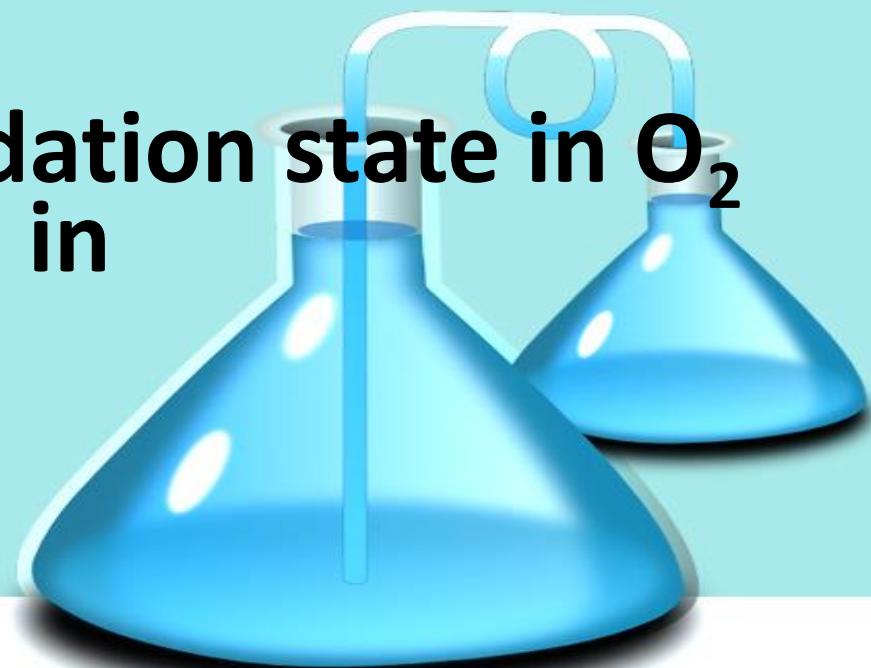


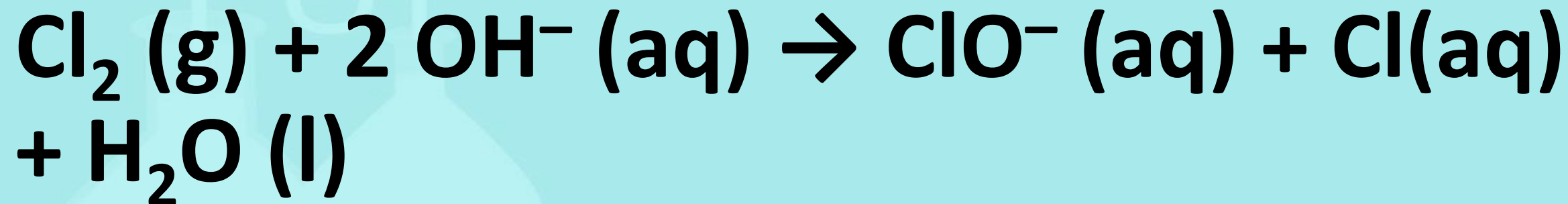
A very familiar example of disproportionation reaction is the decomposition of hydrogen peroxide.



In this reaction, the oxygen of peroxide, which is present

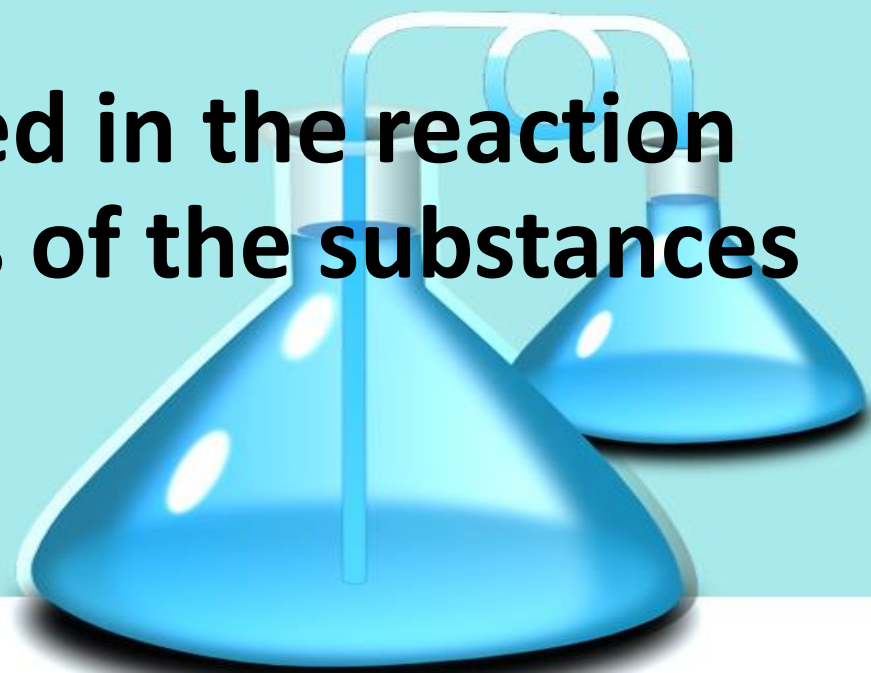
in -1 state, is converted to zero oxidation state in O_2 and decreases to -2 oxidation state in H_2O .





The reaction has a practical application as it describes the formation of household bleaching agents.

The hypochlorite ion (ClO^-) formed in the reaction oxidises the colour-bearing stains of the substances to colourless compounds.



- Among the oxoanions of chlorine, ClO_4^- will not disproportionate because in this oxoanion chlorine is present in its highest oxidation state that is, +7.

Presentation by Pratibha Kohli

