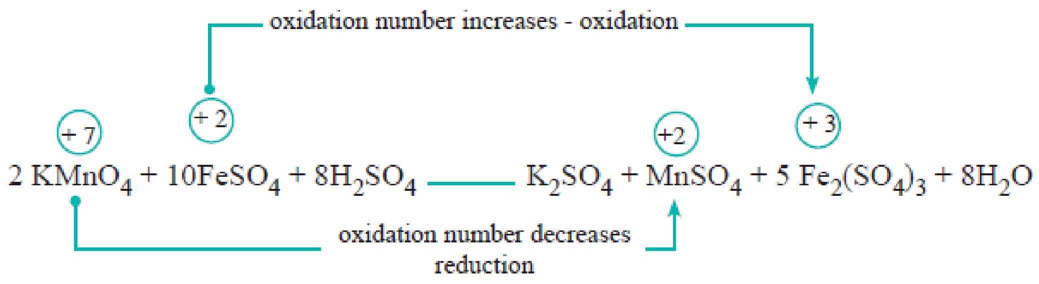
Warm Greetings

Dear students, I think previous uploaded notes are very useful to you. Now we learn about the Redox reactions in terms of oxidation Number and types of redox reaction.

REDOX REACTIONS IN TERMS OF OXIDATION NUMBERS

During redox reactions, the oxidation number of elements changes. A reaction in which oxidation number of the element increases is called oxidation. A reaction in which it decreases is called reduction. Consider the following reaction



In this reaction, manganese in potassium permanganate ($KMnO_4$) favours the oxidation of ferrous sulphate ($FeSO_4$) into ferric sulphate ($Fe_2(SO_4)_3$) by gaining electrons and thereby gets reduced. Such reagents are called oxidising agents or oxidants. Similarly, the reagents which facilitate reduction by releasing electrons and get oxidised are called reducing agents.

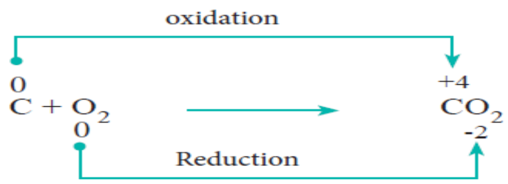
TYPES REDOX REACTION

Redox reactions are classified into the following types.

Combination reactions

Redox reactions in which two substances combine to form a single compound are called combination reaction.

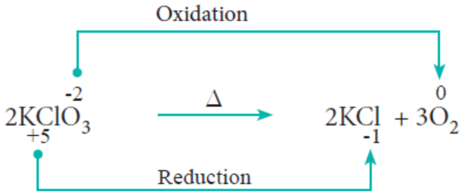
Example



Decomposition reactions

Redox reactions in which a compound breaks down into two or more components are called decomposition reactions. These reactions are opposite to combination reactions. In these reactions, the oxidation number of the different elements in the same substance is changed.

Example

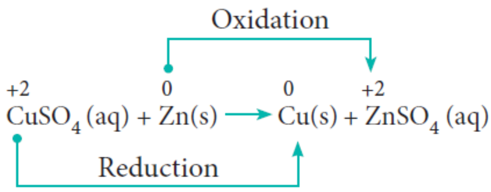


Displacement reactions

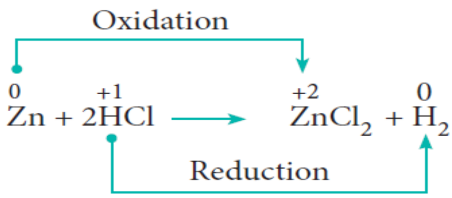
Redox reactions in which an ion (or an atom) in a compound is replaced by an ion (or atom) of another element are called displacement reactions.

Metal displacement reactions

A zinc metal strip in an aqueous copper sulphate solution taken in a beaker. The intensity of blue colour of the solution slowly reduced and finally disappeared. The zinc metal strip became coated with brownish metallic copper.

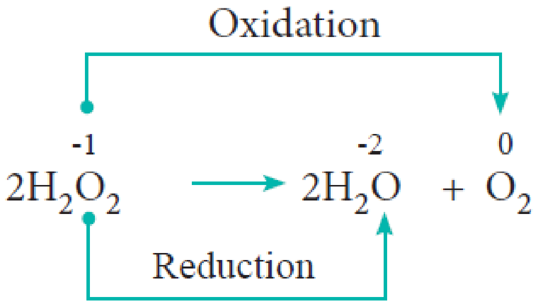


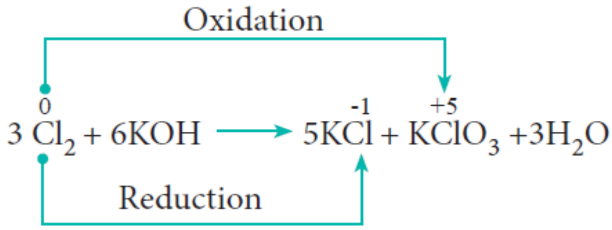
Non-metal displacement



Disproportionation reaction (Auto redox reactions)

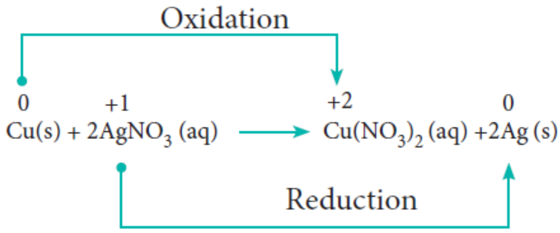
In some redox reactions, the same compound can undergo both oxidation and reduction. In such reactions, the oxidation state of one and the same element is both increased and decreased. These reactions are called disproportionation reactions.





Competitive electron transfer reaction

- In metal displacement reactions, we learnt that zinc replaces copper from copper sulphate solution.
- Place a metallic copper strip in zinc sulphate solution.
- If copper replaces zinc from zinc sulphate solution, Cu^{2+} ions would be released into the solution and the colour of the solution would change to blue.
- But no such change is observed.
- Therefore, we conclude that among zinc and copper, zinc has more tendency to release electrons and copper to accept the electrons.
- The reaction to copper metal and silver nitrate solution.
- Place a strip of metallic copper in sliver nitrate solution taken in a beaker.
- After some time, the solution slowly turns blue.
- This is due to the formation of Cu^{2+} ions, i.e. copper replaces silver from silver nitrate.
- The reaction is,



- It indicates that between copper and silver, copper has the tendency to release electrons and silver to accept electrons.
- The three metals, namely, zinc, copper and silver, the electron releasing tendency is in the following order.

Zinc > Copper > Silver

Balancing (the Equation) of Redox Reactions

The two methods for balancing the equation of redox reactions are as follows.

i) The oxidation number method

ii) Ion-electron method / half reaction method.

Both are based on the same principle:

In oxidation - reduction reactions the total number of electrons donated by the reducing agent is equal to the total number of electrons gained by the oxidising agent.

Oxidation number method

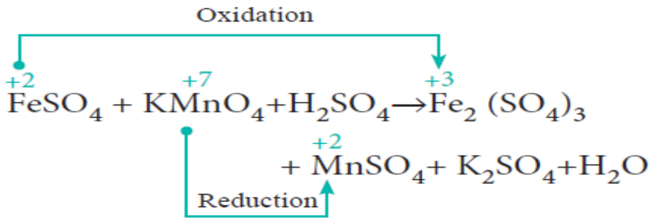
In this method, the number of electrons lost or gained in the reaction is calculated from the Oxidation numbers of elements before and after the reaction. Let us consider the oxidation of ferrous sulphate by potassium permanganate in acid medium.

The unbalanced chemical equation is,

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

Step 1

Using oxidation number concept, identify the reactants (atom) which undergo oxidation and reduction.

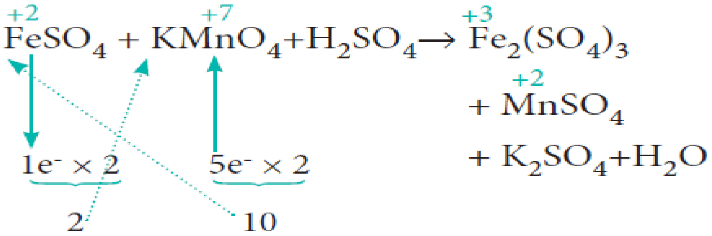


The oxidation number of Mn in $KMnO_4$ changes from +7 to +2 by gaining five electrons.

The oxidation number of Fe in $FeSO_4$ changes from +2 to +3 by losing one electron.

Step 2

Since, the total number of electrons lost is equal to the total number of electrons gained, equate, the number of electrons, by cross multiplication of the respective formula with suitable integers on reactant side as below. Here, the product $Fe_2(SO_4)_3$ contains 2 moles of iron, So, the Coefficients $1e^-$ & $5e^-$ are multiplied by the number '2'.



$$10\ FeSO_4 + 2\ KMnO_4 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 +$$
$$MnSO_4 + K_2SO_4 + H_2O$$

Step 3

Balance the reactant / Product -Oxidised / reduced

Now, based on the reactant side, balance the products (ie oxidized and reduced).The above equation becomes

$$10FeSO_4 + 2KMnO_4 + H_2SO_4 \rightarrow 5Fe_2(SO_4)_3$$
$$+ 2MnSO_4 + K_2SO_4 + H_2O$$

Step 4

Balance the other elements except H and O atoms. In this case, we have to balance K and S atoms but K is balanced automatically.

Reactant Side: 10 'S' atoms (10 $FeSO_4$)

Product Side: 18 'S' atoms

$$5Fe_2(SO_4)_3 + 2MnSO_4 + K_2SO_4$$
$$15S \quad + \quad 2S \quad + \quad 1S \quad = \quad 18S$$

The difference 8-S atoms in reactant side, has to be balanced by multiplying H_2SO_4 by '8' the equation now becomes,

$$10FeSO_4 + 2KMnO_4 + 8H_2SO_4 \rightarrow 5Fe_2(SO_4)_3$$
$$+ 2MnSO_4 + K_2SO_4 + H_2O$$

Step 5

Balancing 'H' and 'O' atoms

Reactant side '16'-H atoms ($8H_2SO_4$ i.e. 8 x 2H = 16 'H')

Product side '2' - H atoms (H_2O i.e. 1 x 2H = 2 'H')

Therefore, multiply H_2O molecules in the product side by '8'

$$10\ FeSO_4 + 2\ KMnO_4 + 8\ H_2SO_4 \rightarrow$$
$$5\ Fe_2(SO_4)_3 + 2\ MnSO_4 + K_2SO_4 + 8H_2O$$

The oxygen atom is automatically balanced. This is the balanced equation

Ion Electron method

This method is used for ionic redox reactions.

Step 1

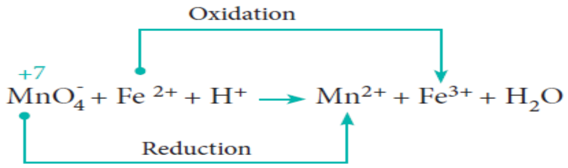
Using oxidation number concept, find out the reactants which undergo oxidation and reduction.

Step 2

Write two separate half equations for oxidation and reduction reaction. Let us consider the same example which we have already discussed in oxidation number method.

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

The ionic form of this reaction is,



The two half reactions are,

$$Fe^{2+} \rightarrow Fe^{3+} + 1e^- \text{ ------------------------ } (1)$$

and

$$MnO_4^- + 5e^- \rightarrow Mn^{2+} \text{ -------------------- } (2)$$

Balance the atoms and charges on both sides of the half reactions.

Equation (1) ⇒ No changes i.e.,

$$Fe^{2+} \rightarrow Fe^{3+} + 1e^- \text{ ------------------------- } (1)$$

Equation (2) ⇒ 4'O' on the reactant side, therefore add $4H_2O$ on the product side, to balance 'H' - add, $8H^+$ in the reactant side

$$MnO_4^- + 5e^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O\text{----}\ (3)$$

Step 3

Equate both half reactions such that the number of electrons lost is equal to number of electrons gained. Addition of two half reactions gives the balanced equation represented by equation (6).

$$(1) \times 5 \Rightarrow \quad 5Fe^{2+} \rightarrow 5Fe^{3+} + \cancel{5e^-} \text{ -------------------- } (4)$$

$$(3) \times 1 \Rightarrow \quad MnO_4^- + \cancel{5e^-} + 8H^+ \rightarrow Mn^{2+} + 4H_2O \text{ -----------------} (5)$$

$$(4) + (5) \quad 5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O \text{-----------} (6)$$

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