CHAPTER I

GENERAL INTRODUCTION

Reaction rate is of great practical interest in both laboratory and industry. The reaction that takes years to become sensibly complete is particularly useful in making its products, but reactions that complete in fraction of seconds include hazardous explosion. Therefore, it becomes necessary for a kinetician to understand the factors controlling the rates. For synthetic purpose, knowledge of reaction mechanism will often allow the reaction condition to be selected for maximum product yields. Finally, it is an area of chemistry where the practical application of theory lacks far behind the realm of experiment, yet at several points theory has had not able successes.

Kinetic studies are receiving much importance in the recent years since they provide us the most powerful method of investigating the detailed reaction mechanisms. It is one of the most intriguing and challenging areas of chemistry, which deals with the mechanisms of reactions. To many chemists the real heart of chemistry is the study of mechanisms. Thus, chemical kinetics can be defined as that branch of chemistry concerned with the study and prediction of time dependent systems. To understand the mechanism of any reaction we must know a reaction as a function of time, the exact positions of all the atoms as the reactants are converted into product molecules. Virtually all information regarding reaction mechanism
comes by inference of indirect evidence. Hence, it is the important job of chemists to device the proper experiments to generate most conclusive evidence.

The experimental part of the subject deals with ways of measuring precisely the rates of reactions at various varying conditions of the experiments. The interpretation of results leads to an understanding of the mechanism of the reaction. The combination of the results of a large number of experiments gives rise to general theories of chemical reactivity. The important steps in any kinetic investigations are; (1) collection of kinetic data (2) establishment of relationships between the rate and reaction mixture composition (3) study of structural effects and (4) interpretation of the collected data to arrive at reaction mechanism.

The award of Nobel prize for the year 1992 to Prof. R. A. Marcus on the “Electron Transfer Reactions” and 1999 Nobel prize to Prof. Ahmed Zewail for discovery of “Femtochemistry” and 2001 Nobel prize to Profs. William Knowles, K. Barry Sharpless and Royji Noyori for their work on “Chirally Catalysed Hydrogenation Reactions” emphasizes the importance of field of reaction kinetics. Electron transfer reactions play a central role in physical, chemical and biological processes. Because of the ubiquity of electron transfer processes, the study of electron transfer reactions, perhaps more so than that of any other area of chemistry is characterized by a strong interplay of theory and experiment. Nonetheless the importance of electron transfer in transition metal redox chemistry has been recognised and more recently it has become increasingly obvious that many
reactions in organic chemistry once thought to be concerted in nature also occur via sequential one electron steps.  

The study of reaction kinetics gives wealth of information about the nature and course of reactions. The order of a reaction can be used to interpret the reaction on molecular level. For example, the chemists by considering the order of a reaction with respect to different reactive species predict the sequence in which bonds break and atoms rearrange during the reaction and hence propose reaction mechanism. 

The kinetic data will be the source of a great deal of detailed insight into the mechanism of a reaction. Although, other types of experimental evidences are also sought for purpose of formulating a reaction mechanism, the study of reaction kinetics generally forms the backbone of a thorough mechanistic investigation. Finally, as an area of pure science in itself, the study of rates and mechanism is one of rich varieties, concerned with the chemistry of every element and full of experimental challenges. Thus, chemical kinetics can be defined as the branch of physical chemistry devoted to the laws of chemical reactions in time. 

The work of Henry Taube in redox systems unequivocally demonstrated the transport of electron from reductant to oxidant. This discovery certainly added many important features in the syntheses of coordination complexes and organometallics. It is such a subject, which has manifestations in almost all walks of life. As a result, oxidation-reduction reaction needs at least two reactants, one capable of gaining electrons (oxidant) and the other capable of losing electrons.
(reductant), i.e., a reducing agent (reductant) by losing electrons, gets oxidised and an oxidising agent (oxidant), by gaining the electrons, gets reduced.

The reaction thus consists of transfer of electrons from a reducing agent to an oxidising agent, so that there cannot be an oxidation without concomitant, reduction. Such cases where oxidation and reduction are involved are commonly termed as redox reactions; and such a redox system involves a redox potential. It follows automatically, that any oxidation-reduction reaction must involve two redox couples that differ in their affinity for electrons. This affinity of atoms for electrons is conferred to them by their particular atomic structure and it is expressed in terms of reduction potential or redox potential. Reduction potential of a redox couple is its tendency to get reduced from one oxidation state to another oxidation state. Redox reactions are also defined as involving changes in oxidation states or oxidation numbers. In a redox reaction, the oxidation number of the oxidant decreases and that of the reductant increases. The transfer of electrons is a book-keeping device for effecting the changes in oxidation states and for balancing the equations.

Particularly in aqueous solutions, it is usually possible to imagine atom or group transfer, rather than electron transfer, as occurring in a redox reaction. For example, iron (II) ion may act as a reducing agent by transferring a hydrogen atom from its hydration shell to a substrate:

\[ \text{Fe(H}_2\text{O)}_6^{2+} + R' \rightarrow \text{Fe(H}_2\text{O)}_5\text{OH}^{2+} + RH \]

Iron(III) ion may act as an oxidising agent by transferring hydroxyl radical to a substrate:

\[ \text{Fe(H}_2\text{O)}_6^{3+} + R' \rightarrow \text{Fe(H}_2\text{O)}_5^{2+} + \text{H}^+ + \text{ROH} \]
In general, transfer of a positive group or atom is equivalent to the transfer of electrons and transfer of a negative group or atoms is equivalent to the taking up of electrons. The problem, then, in studying the mechanism of an oxidation–reduction reaction, is to find out whether atom transfer or electron transfer occurs, which atoms are transferred, and what intermediates, stable or unstable species are formed. A complete study would include a detailed picture of the transition state for all steps involved. Not only the composition but also the geometry of the transition state is desired.

**Oxidation-Reduction in Inorganic Reactions**

Two general classes of transition states emerge for redox reactions involving metal complexes, the so-called **outer-sphere** and **inner-sphere types**. In the first of these, the inner coordination shells of both the metal ions are intact in the transition state. In the second case, the two metal ions are connected through a bridging ligand common to both the coordination shells. From the **Franck-Condon principle**, it follows that before electron transfer between two ions is possible, the energy of the electron must be the same in the two sites. There must also be sufficient orbital overlap between the two sites to provide for a reasonable probability of a transfer.

In the case of reaction of **outer-sphere** type, the electron must wait for appropriate fluctuation of the ions and their coordination spheres, without in any way being able to influence the changes required in the site to which it will be transferred. For reactions of inner-sphere type, an activated complex of the type \( L \rightarrow M^{n+1} - X-M^n W_x \) may be formed, where \( L, X \) and \( W \) are ligands. In arriving at a
configuration having $X$ as a bridging group, either $M^{n+1}$ or $M^n$ (or both) has undergone substitution in the first coordination sphere. Delocalisation of an electron over two sites will lower the energy needed for the formation of the activated complex and the electron will be able to affect the energy required to produce a fluctuation at the site to which it is going to be transferred.

Presumably the reaction is consummated by the fluctuation which causes the separation of $M^{n+1}$ and $M^n$ and may well require other changes in the coordination spheres as well. A feature of the bridged activated complex is that the bridging group may move from the oxidising agent to the reducing agent.

The most conclusive evidence for the mechanism comes from isolation of the products, which is possible only if the appropriate solvolysis reaction occurs slowly. However, the reaction of Cr(II) and some of the reactions of V(II) and Fe(II) are convincingly characterized as following the inner-sphere mechanism. In the latter case, the immediate V(III)-X or Fe(III)-X products have often been detected spectrophotometrically, prior to their dissociation$^{6-8}$. Therefore, probably the inner-sphere mechanism is the most favourable provided, (a) a suitable ligand is available in the coordination sphere of the oxidising complex and (b) substitution of a water molecule in the reducing agent occurs sufficiently rapidly to permit the transition state $M-X-M'$ to form rapidly$^{9a}$. The most common instances of outer-sphere mechanism are found when nature of the coordination sphere of one reductant of other contains ligands unsuited for bridging or when exchange occurs too slowly to permit bridge formation. Among such reactions$^{9b}$ are Co(NH$_3$)$_6^{3+}$.
oxidations and Cr(bipy)$_3^{2+}$ reductions; certain reactions of V(H$_2$O)$_6^{2+}$, which undergo substitution more slowly than most divalent ions; electron exchange of Fe(C$_3$H$_5$)$_2$ and Fe(C$_3$H$_5$)$_2^{2+}$; and reactions such as electron exchange of Fe(CN)$_6^{3-}$ and Fe(CN)$_6^{4-}$.

**Oxidation-Reduction in Organic Reactions**

The oxidation-reduction concepts, however, are not so clearly applicable in organic chemistry, for when carbon compounds are oxidised their component atoms are very seldom deprived of their surrounding complete electron shells. Covalent bond fission is an essential feature of organic reactions and it can be effected by two different pathways$^{10}$, viz., "Homolytic reactions" in which electron pairs are symmetrically disrupted and "Heterolytic reactions" in which electron pairs are transferred from one molecule to another as an undivided entity. Electron removal by these two pathways has clearly distinguishable characteristics.

In homolytic oxidations electrons are removed singly from organic molecules by active atoms such as chlorine, or by active free radicals. Though, molecules containing unshared electrons can be oxidised in this way, homolytic oxidations usually involve the removal from an organic molecule of one electron together with a hydrogen atom e.g.,

$$\text{R}_3\text{C-H} + \text{Cl}^- \rightarrow \text{R}_3\text{C}^- + \text{HCl}$$

The initial organic product necessarily has unpaired electron and so must perforce undergo a reaction of similar type e.g.,

$$\text{R}_3\text{C}^- + \text{Cl-Cl} \rightarrow \text{R}_3\text{C-Cl} + \text{Cl}^-$$
or must combine with another free radical before stable entities alone result. Thus, chain reactions, dimerisations or disproportionations e.g.,

\[ 2 \text{C}_2\text{H}_5^* \rightarrow \text{C}_2\text{H}_6 + \text{C}_2\text{H}_4 \]

are typical homolytic reactions.

All homolytic electron transfer reactions, such as above require very much less activation energy than needed for direct breakage of a covalence. Hence, homolytic oxidations, when once started, proceed very rapidly indeed. The traces of free radicals required for the initiation of homolytic oxidation may be formed by thermal dissociation of molecules which have weak covalences, by disruptions of molecules by exposure to radiant energy, high energy particles (x-particles) or electrons (β-particles) or by single electron transfer, from ions of transition elements which can have incomplete inner electron shells (d \rightarrow). Consequently oxidations which are demonstrably prone to catalysis in these ways are homolytic in type.

Heterolytic oxidations involve the attack on organic compounds of electrophilic reagents, which can, by a single process gain control of a further electron pair. Heterolytic oxidants therefore attack the exposed electron pairs of atom such as oxygen, nitrogen or sulphur or the loosely held π-electrons of olefins, rather than buried electron pairs of C-H, O-H or N-H bonds. Again, heterolytic reactions yield stable molecular or ionic products in one or at most two consecutive stages and very seldom lead to chain reactions. On the other hand, they usually require higher activation energy than homolytic reactions and so tend to be slower processes. Use of catalyst lowers the activation energy in such cases.
heterolytic reactions are subjected to catalyst, the mechanism of catalysis is largely
different from the catalysis involved in the case of homolytic reactions.

Probable ways of electron transfer reactions

Oxidation-Reduction reaction may involve one or more electron transfer. Depending upon the number of electrons transferred between oxidant and reductant, the reaction may proceed in one or more steps. Transition metals such as iron and cobalt and several others usually exhibit stable oxidation states differing by one electron and react with each other through one equivalent steps. However, the stable oxidation states in post transition elements such as arsenic, antimony etc., differ by two electrons. Thus, on the basis of their pattern of reactivity, the reactions of these elements are classified into two main categories \(^4,11-16\),

Complementary and Non-complementary reactions

Complementary Reactions

Complementary reactions are those in which oxidant and reductant both undergo two-equivalent changes or one-equivalent change. These reactions generally obey a bimolecular rate equation and the electron transfer can take place in a single step or in two step of one electron each. In Tl(I) - Tl(III) exchange reactions, electrons are transferred in a single step as the formation of the intermediate, Tl(II), is not detected. On the other hand the oxidations of As(III) \(^{17}\) and Sb(III) \(^{18}\) by Tl(III) are other reactions where no evidence for the formation of Tl(II) has been obtained. The evidence obtained for the formation of As(IV) in pulse radiolytic studies indirectly supports one-equivalent steps in As(III)
oxidations. However, such an evidence in the reaction of As(III)-Tl(III) has not been observed. It is assumed that As(IV) is formed in the reaction (i) then it is very much likely that the two intermediates Tl(II) and As(IV) react with each other before they can diffuse out of the solvent cage in which they are found.

\[
\text{As(III)} + \text{Tl(III)} \rightleftharpoons \text{Tl(II)} + \text{As(IV)} \tag{i}
\]

A large number of complementary reactions have been explained by assuming the formation of bridged activated complex between oxidant and reductant for the facile transfer of electron through the bridging ligand.

**Non-Complementary Reactions**

Non-complementary reactions are those in which oxidant and reductant undergo unequal equivalent changes such as one-equivalent oxidant interacts with two-equivalent reductant and two-equivalent oxidant interacts with one-equivalent reductant. There are a number of possibilities of electron transfer in non-complementary reactions and these are related to the nature of both oxidant and reductant.

**Multi equivalent reactions**

Oxidising agents such as Cr(VI) and Mn(VII) undergo net changes of 3 and 5 units in oxidation number respectively during their reactions in acidic solution. For the most part, these reactions occur by one or two electron steps, with the necessary intervention of unstable intermediate oxidation states of Cr or Mn. The reactions of Cr(VI) with transition metal complexes generally proceed by sequential one-
electron step, but with post transition metal ions and with non-metallic compounds, two electron steps appear to be preferred.

Electron transfer reactions are found to be governed by two classical principles

(a) Michaelis principle of compulsory univalent oxidation steps

(b) Shaffer's principle of equivalent change

Michaelis hypothesis involves the principle that an oxidation-reduction reaction takes place in one or more successive single electron transfer steps. This principle evolved from a considerations of restricted field of redox reactions, of which the oxidation of hydroquinones to quinones through semiquinone intermediate is typical and is now generally recognised as being without universal validity. Apart from the reactions involving metal ions, many two equivalent redox reactions are now known which proceed in one step through the transfer of hydride ion or an oxygen atom.

\[ \text{e.g., } \text{NO}_2^- + \text{OCl}^- \rightarrow \text{NO}_3^- + \text{Cl}^- \]

The second principle refers to the observation that non-complementary reactions are often slow compared with complementary one's. Examples are the slow reduction of Tl(III) by Fe(II) or Ce(IV) by Tl(I) as compared to the rapid reduction of Tl(III) by Sn(II) and Ce(IV) by Fe(II).

One of the implications of the comparison on which the principle of equivalence change is based is that reactions between 2-equivalent oxidants and 2-equivalent reductants occur by a concerted 2-equivalent step. This may well be the
case for reactions such as the exchange $\text{Tl(I)} \rightarrow \text{Tl(II)}$ \cite{24,26}. The observations expressed by Shaffer, for non-complementary reactions, is based on the low probability of termolecular mechanisms as one possibility or the formation of the unstable valence states as the other possibility\cite{15}. Another expected feature of these reactions will be the frequent occurrence of catalysis since the normal paths will be slow.

**Unstable oxidation states**

The formation of unstable oxidation states during the course of non-complementary reactions has been now anticipated in a number of such reactions with sufficient proofs. For example, the reductions of $\text{Tl(III)}$ by $\text{Fe(II)}$\cite{27,29}, $\text{V(III)}$ or $\text{V(IV)}$\cite{30,33} and $\text{Cr(VI)}$ by $\text{Tl(I)}$\cite{34} can only be explained through the formation of unstable $\text{Tl (II)}$ species. Similar unstable oxidation states have been observed in other studies\cite{32,33}. The interconversions between $\text{Cr(III)}$ and $\text{Cr(VI)}$ always appear to involve the unstable states, $\text{Cr(IV)}$ and $\text{Cr(V)}$.

In a classic study, King and Tong\cite{35} have worked out the details of the redox reactions between $\text{Ce(IV)}$ and $\text{Cr(III)}$ in aqueous sulphuric acid. The rate law was found to be as in equation (ii) which is very reasonably explained by the mechanism involving steps of equation (iia) to (iic).

$$\text{Rate} = k \frac{K[\text{Ce(IV)}]^2[\text{Cr(III)}]/[\text{Ce(III)}]}$$ \hspace{1cm} (ii)

The first step is a rapid equilibrium, and a second step, the interconversion of $\text{Cr(IV)}$ to $\text{Cr(V)}$ is rate determining.
Excellent support comes from the study of related reactions such as the oxidation of vanadyl ion by acid chromate ion\textsuperscript{36} (HCrO$_4$) and the analytical important oxidation of ferrous ion by acid chromate\textsuperscript{37}. It is significant that in the above example the change over from Cr(IV) to Cr(V) or vice versa, is rate determining. This may be related to the likelihood that, at this stage, a change in coordination number from 6 to 4 occurs\textsuperscript{35}. The stoichiometric examination of the reaction between acid chromate and Fe(II) or VO$^{2+}$ in presence of iodide leads to the conclusion that the rapid oxidation of iodide\textsuperscript{38} is because of its reaction with Cr(V).

A related phenomenon is seen in the oxidation of As(III) by peroxodisulphate ion. This slow reaction is accelerated or oxidation is induced by adding Fe(II) and the reaction is postulated to go by the formation of As(IV) from the reaction (iiia, iiib).

\[
\begin{align*}
\text{Fe(II) + S}_2\text{O}_8^2^- & \longrightarrow \text{Fe(III) + SO}_4^{2-} + \text{SO}_4^- & \text{(iiia)} \\
\text{As(III) + SO}_4^- & \longrightarrow \text{As(IV) + SO}_4^{2-} & \text{(iiib)}
\end{align*}
\]

As mentioned earlier, another feature of these unstable oxidation states is the frequent occurrence of catalysis. A number of studies of the catalysis by platinum metals of oxidation reactions have been made\textsuperscript{39}. The catalysis by Ag(I), Cu(II),
Mn(II) and Cr(III) in oxidation-reduction reactions are also found to occur through formation of unstable oxidation states\textsuperscript{40}.

Active Species

If a particular substance (oxidant, reductant or catalyst) is capable of existence in several forms in aqueous solution, all the species existing may not be active. Those species, which are involved in a slow step, will influence the reaction. The reaction conditions will determine the nature of the active species.

Moore and Anderson\textsuperscript{41} have reported that three or four species of Ce(IV) exist in sulphuric acid depending on its strength and the main active species in dilute solutions is Ce(SO$_4$)$_2^{2+}$. Hardwick and Robertson\textsuperscript{42} have give the following equilibria amongst various species of Ce(IV) in sulphuric acid solutions. The concentration of H\textsuperscript{+}, HSO$_4^{-}$ and SO$_4^{2-}$ would determine the nature of the active species.

\[
\begin{align*}
\text{Ce}^{4+} + \text{HSO}_4^- & \rightleftharpoons \text{CeSO}_4^{2+} + \text{H}^+ \\
\text{CeSO}_4^{2+} + \text{HSO}_4^- & \rightleftharpoons \text{Ce(SO}_4)_2 + \text{H}^+ \\
\text{Ce(SO}_4)_2 + \text{HSO}_4^- & \rightleftharpoons \text{H}_2\text{Ce(SO}_4)_3^- \\
\text{H}_2\text{Ce(SO}_4)_3^- + \text{HSO}_4^- + \text{H}^+ & \rightleftharpoons \text{H}_3\text{Ce(SO}_4)_4^- 
\end{align*}
\]

To formulate the reaction rate as a function of species concentration, therefore, requires knowledge of the existence of such equilibria and generally speaking, the knowledge of determination of one or more equilibrium constants. The distinction between the species and laboratory concentrations is particularly
critical in the cases of partially displaced equilibria, because the rate equations, if cast in the form of reaction rate as a function of laboratory concentration, quite often resemble the equations applicable to different mechanisms.

**Effect of ions on the rate**

The rate of redox reactions of complex ions is highly sensitive to the presence of other ions in the solution. Usually, added anions have the greater influence on the reaction of two cations each other and added cations influence the rate of two anions with each other. Large effects of anions on the rate are usually attributed to the formation of bridged species if labile complexes are involved.

In the redox reaction between two cations, if the reducing agent is complexed first, it will be more stabilized in the oxidized form and hence the reaction is speeded up in the presence of anion. Also a negative ion should enable to positive ions to approach each other more easily. This will be more effective if the negative ion is between the two cations in the transition state.

If the oxidising agent is complexed first the anion may stabilize it and slow down its rate of reaction. For example, the rate of reactions of Ce(IV) oxidations are reduced strongly by forming sulphate complexes\(^{43}\) in presence of sulphate. This shows, in agreement with the theory that the coulombic repulsions between reactants do not play a dominant role. The conclusions to be drawn are not that electrostatic factors do not exist, but that other factors are of greater importance. Some significant effects of chloride ions are particularly noticeable in redox
reactions. The redox reactions in which the reductant is able to form complexes with chloride are Tl(I), Pd(II), Sn(II) and Sb(III).

**Catalysis**

Any substance, other than reactants which influences the rate of chemical reaction and itself remains unchanged chemically at the end, is called a catalyst. The phenomenon of rate alteration is designated as catalysis. Catalysts influence the reactions by changing the reaction path. Such catalytic influences arise as consequences of lowering of the energy of activation.

Catalyst development and synthesis have become a wide spread research field because of the increasing global demand for better systems in chemical industry. Today, the chemical industry has been blamed for producing environmentally hazardous substances, which cause acid rain, a reduction of stratospheric ozone levels and so on. Many industrial processes have become burden on the environment, and, therefore must be essentially replaced by more eco-friendly or compatible processes in addition to this, disposing of byproducts or converting them into environmentally non-hazardous substances consume much energy. The successful exploitations of a material as a catalyst will give value-added products with improved yields and also eliminate or minimise environmental pollutants.

In solution involving inorganic oxidations, the catalysts are ions having unstable oxidation states. This case is a particular example of homogeneous catalysis where the catalyst present in the same phase as a part of reactants.
Though, the mechanism of catalysis depends on the nature of the substrate, oxidant and other experimental conditions, it has been shown that metal ion acts as catalyst by one of the following paths\textsuperscript{44}.

(a) The catalyst is first oxidised by the oxidant to its higher oxidation states which in a subsequent step oxidises the substrate.

(b) The catalyst forms a complex with oxidant, the complex then oxidises the substrate either in the rate determining or in a fast step.

(c) The catalyst itself first oxidises the substrate in a slow step and the reduced form of the catalyst is oxidised by the oxidant in a fast step.

(d) The catalyst traps the radical produced as intermediates and oxidises them at faster rate.

Osmium(VIII) is a case in point as it has been utilized\textsuperscript{10} as catalyst for hexacyanoferrate(III) oxidation of organic and inorganic substrates. The catalysis is understood to be due to the intervention of intermediate oxidation states of osmium. The detailed picture of its catalytic action has been reviewed by Upadhaya et al\textsuperscript{45}.

Often, very small quantities of catalysts like Os(VIII)\textsuperscript{46}, Pd(II)\textsuperscript{47}, Cr(III)\textsuperscript{48}, Ru(III)\textsuperscript{49}, V(V)\textsuperscript{50} etc., causes appreciable rates accelerations of particular reactions\textsuperscript{39,40}.

**Present Work**

In the present thesis, some redox reactions in alkaline or acid medium as detailed below have been studied. Reactions were followed conveniently by spectrophotometer in the uv-vis. region. The details of such studies are given below.
The present investigation includes some uncatalysed and catalysed oxidation reactions respectively. The thesis is divided into seven chapters comprising the following.

1. **General Introduction**

   This chapter introduces about the kinetics and mechanism of reactions.

2. **Kinetics and Mechanism of Oxidation of L-Leucine by alkaline Diperiodatonickelate(IV) - A free radical intervention, deamination and decarboxylation**

   In this chapter the kinetics of oxidation of L-Leucine by alkaline diperiodatonickelate(IV) was studied spectrophotometrically using a rapid kinetic accessory. The reaction is first order with respect to [DPN] and is an apparent less than unit order, each in [L-leucine] and [alkali] under the experimental conditions. However, the order in [L-leucine] and [alkali] changes from first order to zero order as the concentrations change from lower to higher concentrations respectively. Addition of periodate has no effect on the reaction rate. A mechanism involving the deprotonated diperiodatonickelate(IV) (DPN) as the reactive oxidant species is proposed. The reaction constants involved in the mechanism were evaluated. The isokinetic temperature was determined. The activation parameters with respect to the slow step of scheme were evaluated and discussed.

3. **Kinetics and Mechanism of Oxidation of L-Proline by Heptavalent Manganese: A free radical intervention and decarboxylation (Stopped flow technique)**
The kinetics of oxidation of L-proline by permanganate in alkaline medium was studied spectrophotometrically which is connected to a rapid kinetic accessory. The reaction is first order with respect to [MnO₄⁻] and is an apparent less than unit order, each in [L-proline] and [alkali] under the experimental conditions. Reaction rate increases with increase in ionic strength and decrease in solvent polarity of the medium. Addition of reaction products has no effect on the reaction rate. A mechanism involving the formation of a complex between the oxidant and substrate has been proposed. The reaction constants involved in the mechanism were evaluated. There is a good agreement between the observed and calculated rate constants under varying experimental conditions. The activation parameters with respect to the slow step of scheme were evaluated.

4. **Kinetic and Mechanistic study of Oxidative degradation of Isonicotinate ion by Diperiodatonicelate (IV) in Aqueous Alkaline Medium**

The kinetics of oxidation of isonicotinate ion by alkaline diperiodatonicelate(IV) was studied spectrophotometrically. The reaction is first order with respect to [DPN] and is an apparent less than unit order, each in [isonicotinate ion] and [alkali] under the experimental conditions. However, the order in [isonicotinate ion] and [alkali] changes from first order to zero order as the concentrations change from lower to higher concentrations respectively. Reaction rate increases with increase in ionic strength and decrease in solvent polarity of the medium. Addition of periodate has no effect on the reaction rate. A mechanism the
involving deprotonated diperiodatonickelate(IV) (DPN) as the reactive oxidant species has been proposed.

5. Kinetics and Mechanistic study of Oxidative Decarboxylation of Aspirin by Diperiodatocuprate (III) in Aqueous Alkaline Medium

The reaction is first order with respect to [DPC] and is an apparent less than unit order, each in [Aspirin] and [alkali] under the experimental conditions. However, the order in [Aspirin] and [alkali] changes from first order to zero order as the concentrations change from lower to higher concentrations respectively. Addition of periodate has no effect on the reaction rate. A mechanism involving the deprotonated diperiodatocuprate (DPC) as the reactive oxidant species has been proposed. The reaction follows a 1:4 stoichiometry as given below,

\[
\text{C}-\text{CH}_3 + 4\text{Cu(III)} + 2\text{OH}^- \quad \rightarrow \quad \text{CH}_3\text{COOH} + 2\text{H}^+ \quad (1)
\]

6. Kinetic, mechanistic and spectral studies for the oxidation of Atenolol by alkaline hexacyanoferrate(III)

The kinetics of oxidation of atenolol (ATN) 4-(2-hydroxy-3-isopropylaminopropoxy) phenylacetamide by hexacyanoferrate (III) in alkaline medium was studied spectrophotometrically. The reaction showed first order kinetics in hexacyanoferrate (III) and an order of less than unity in both [ATN] and alkali concentrations. The rate of reaction increases with increase in alkali
concentration. Increasing ionic strength increases the rate but the dielectric constant of the medium has no significant effect on the rate of the reaction. A retarding effect was observed by one of the products i.e. hexacyanoferrate (II) (HCF (II)). The main products were identified by I.R., N.M.R., fluorimetric and mass spectral studies. A mechanism involving the formation of a complex between atenolol and hexacyanoferrate (III) has been proposed. The reaction constants involved in the mechanism are evaluated. There is a good agreement between the observed and calculated rate constants under different experimental conditions. Investigations at different temperatures allowed the determination of the activation parameters with respect to the slow step of the proposed mechanism and mechanism.

7. Kinetics and Mechanism of ruthenium(III) catalysed Oxidation of Gabapentin(GP) (Neurontin) by Heptavalent Manganese: A free radical intervention and decarboxylation

The kinetics of ruthenium(III) catalysed oxidation of Gabapentin (1-(aminomethyl) cyclohexanacetic acid)(GP) by permanganate in alkaline medium was studied spectrophotometrically. The reaction is first order with respect to $[\text{MnO}_4^-]$ and [catalyst] with an apparent less than unit order in [Gabapentin] and [alkali] under the experimental conditions. A mechanism involving the formation of a complex between the active species of Ru(III) and substrate has been proposed. The rate law is given by

\[
\text{Rate} = \frac{d[\text{MnO}_4^-]}{dt} = \frac{k K_1 K_2 [\text{GP}][\text{MnO}_4^-][\text{OH}^-][\text{Ru(III)}]}{1 + K_1 K_2 [\text{GP}][\text{OH}^-] + K_1 [\text{OH}^-]}
\]

21
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