Kinetics of oxidative degradation of L(+)lysine by alkaline permanganate - A mechanistic approach

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The kinetics of oxidation of L(+)lysine by alkaline permanganate follows the rate law

\[ \frac{-d[MnO_4^-]}{dt} = \frac{kK_1K_2[L(+)lysine][OH]}{1 + K_1[OH] + K_2[L(+)lysine][OH]} \]

where \(K_1\) and \(K_2\) are the equilibrium constant of permanganate-alkali complex and formation constant for the complex between L(+)lysine and permanganate respectively and \(k\) is the decomposition rate constant. The oxidation of L(+)lysine proceeds via the formation of a complex between L(+)lysine and permanganate which decomposes in a rate limiting step followed by a fast reaction between the free radical of L(+)lysine and another molecule of permanganate to give the products. The activation parameters are computed with respect to the slow step of the mechanism.

Permanganate ion oxidises a greater variety of substrates, it finds extensive applications in organic synthesis especially after the advent of phase transfer catalysis. Kinetic studies on unsaturated acids both in aqueous and non-aqueous media, have been reported. The manganese intermediates are relatively easy to identify provided they have sufficiently long life times, and oxidation states of the intermediates permit useful conclusions concerning the reaction mechanism, including the nature of intermediates. If \([OH] > 0.1\) mol dm\(^{-3}\), permanganate will be reduced only to manganate in the first step and owing to its much lower reactivity, further reaction to manganate will be much slower. Thus the process Mn(VII) to Mn(IV) can be divided into a number of partial steps and examined separately. The MnO\(_4\)\(^-\) appears only after long time i.e., after the complete consumption of MnO\(_2\)\(^-\). In a strongly alkaline medium the stable reduction product of permanganate ion is manganate, MnO\(_2\)\(^-\). No mechanistic information is available to distinguish between a direct one-electron reduction to Mn(VI) (scheme 1)

\[ \begin{align*}
\text{Mn(VII)} + S & \rightarrow \text{Mn(VI)} + S^* + k_f \\
\text{Mn(VII)} + S^* & \rightarrow \text{Mn(VI)} + \text{Products} + k_2
\end{align*} \]

Where \(S = \text{Substrate}, k_2 \gg k_f\)

Scheme 1

and a mechanism in which a hypomanganate is formed in a two electron step followed by a rapid reaction (Scheme 2).

\[ \begin{align*}
\text{Mn(VII)} + S & \rightarrow \text{Mn(V)} + \text{Products} + k_f \\
\text{Mn(VII)} + \text{Mn(V)} & \rightarrow 2 \text{Mn(VI)} + k_4
\end{align*} \]

where \(S = \text{Substrate}, k_4 \gg k_f\)

Scheme 2

Amino acids have been oxidised by a variety of oxidising agents. Different products have been identified by different workers on the oxidation of amino acids. The study of amino acids becomes important because of their biological significance and selectivity towards the oxidants to yield different products. Although a variety of organic and inorganic substrates are oxidised by permanganate in aqueous alkaline medium, there are only a few reports on the oxidation of amino acids by aqueous alkaline permanganate. However, no work has been reported on the oxidation of L(+)lysine. The title study has been undertaken in order to understand the redox chemistry of permanganate in such a medium.

Experimental

Stock solutions of L(+)lysine (s.d.lime-chem) and potassium permanganate (BDH) were prepared by dissolving the appropriate amounts of respective samples in doubly distilled water, the solution of permanganate was standardised against oxalic acid. Potassium mangan-
Table 1: Effect of \([\text{L(+)lysine}], [\text{Mn(VII)}] \) and \([\text{OH}^-] \) on oxidation of L(+)lysine by permanganate in aqueous alkaline medium at 25°C and \(I = 0.6 \text{ mol dm}^{-3}\).

<table>
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<th>([\text{MnO}_4^-] \times 10^4 ) mol dm(^{-3})</th>
<th>([\text{L(+)lys}] \times 10^4 ) mol dm(^{-3})</th>
<th>([\text{OH}^-] ) mol dm(^{-3})</th>
<th>(k_\text{obs} \times 10^{1}, \text{s}^{-1})</th>
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ate solutions were prepared as described by Carrington Symons and standardised spectrophotometrically employing 1 cm quartz cell at 608 nm (ε = 1530 ± 20 dm\(^3\) mol\(^{-1}\) cm\(^{-1}\)).

All other reagents were of analytical grade. NaOH and NaClO\(_4\) were used to maintain required alkalinity and ionic strength respectively.

**Kinetic studies**

All kinetic measurements were carried out under pseudo-first order conditions where \([\text{L(+)lysine}]\) was at least in ten fold excess over permanganate ion at a constant ionic strength of 0.6 mol dm\(^{-3}\). The reaction was initiated by mixing thermally equilibrated solution of \(\text{MnO}_4^-\) and \(\text{L(+)lysine}\) which also contained required quantities of \(\text{NaOH}\) and \(\text{NaClO}_4\) to maintain the required alkalinity and ionic strength respectively. The temperature was uniformly maintained at 25 ± 0.1°C. The course of reaction was followed by monitoring the decrease in the absorbance of \(\text{MnO}_4^-\) in 1 cm quartz cell of Hitachi 150-20 spectrophotometer at its absorption maximum 526 nm as a function of time. Preliminarily it was verified that there was negligible interference from other reagents at this wavelength. The application of Beer’s law for permanganate at 526 nm had earlier been verified giving \(ε = 2083 ± 50 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}\) (literature value 2200). The first order rate constant, \(k_\text{obs}\) was evaluated from plots of \(\log \left[ A_t / A_\infty \right] \) versus time, where \(A\) refers to absorbance and \(A_\infty\) at infinite time. The first order plots in almost all cases were linear up to 85% completion of the reaction and \(k_\text{obs}\) were reproducible within ± 5%.

In the course of measurement the colour of the solution changes from violet to blue and further to green. The spectrum of the green solution was identical to that of \(\text{MnO}_4^2-\). It is evident that the blue colour originated from violet of permanganate and green from manganate excluding the accumulation of hypomanganate.

The effect of dissolved oxygen on the rate of reaction was checked by preparing the reaction mixture and following the reaction in an atmosphere of nitrogen. No significant difference was observed between the results obtained under the nitrogen and in presence of air. Added carbonates had no effect on the reaction rate. Fresh solutions were used while conducting the experiments.

**Results and discussion**

**Stoichiometry**

The reaction mixture containing the excess permanganate over \(\text{L(+)lysine}\) was mixed in the presence of 0.3 mol dm\(^{-3}\) NaOH adjusted to constant ionic strength of 0.60 mol dm\(^{-3}\). After the elapse of reaction time, solid KI was added following acidification by \(\text{H}_2\text{SO}_4\) (10%) then remaining permanganate was titrated against standard sodium thiosulphate. The results indicated that two moles of \(\text{MnO}_4^-\) consumed one mole of \(\text{L(+)lysine}\) as given by Eq.(1).
R-CH(NH₂)COOH + 2 MnO₄⁻ + 2 OH⁻ → R-CHO + NH₃ + 2 MnO₄²⁻ + CO₂
R = H₂N-CH₂-CH₂-CH₂-CH₂-

The main reaction products were identified as the corresponding aldehyde by spot test, ammonia by Nessler's reagent and manganate spectrophotometrically by comparing its spectrum with that of the authentic sample. The corresponding aldehyde was confirmed by preparing its 2,4-DNP derivative. Such products were also obtained in previous work. It was observed that the aldehyde did not undergo further oxidation under the present kinetic conditions.

**Reaction order**

The reaction orders were determined from the slopes of log k<sub>obs</sub> versus log concentration plots by varying the concentration of oxidant, reductant and alkali in turn while keeping concentrations of others constant.

The oxidant, [potassium permanganate] was varied in the range of 5.0 x 10⁻⁴ to 5.0 x 10⁻³ mol dm⁻³ and the linearity of plots of log[MnO₄⁻] versus time indicates the order in [MnO₄⁻] as unity. This was also confirmed by varying [MnO₄⁻] which did not show any change in pseudo-first order rate constants, k<sub>obs</sub> (Table 1). The substrate, L(+)-lysine was varied in the concentration range of 5.0 x 10⁻⁴ to 5.0 x 10⁻³ mol dm⁻³ at 25°C keeping all other [reactants] constant. The order in [L(+)-lysine] was found to be less than unity (Table 1).

The effect of alkali on the reaction was studied at constant [L(+)-lysine] and [KMnO₄] keeping constant ionic strength of 0.60 mol dm⁻³ at 25°C. The rate constant increased with the increase in [OH⁻] (Table 1).

**Effect of ionic strength and solvent polarity**

The effect of ionic strength was studied by varying the [sodium perchlorate] in the reaction mixture. The ionic strength of the reaction medium was varied from 0.30 to 2.0 mol dm⁻³ at constant [permanganate], [L(+)-lysine] and [alkali]. It was found that the rate constant enhanced with increase in [NaClO₄] and the plots of log k<sub>obs</sub> versus 1/D was linear with positive slope (Fig. 1).

The relative permittivity effect was studied by varying the t-butanol content keeping the other conditions fixed. Attempts to measure the relative permittivity was not successful. However, they were computed from the values of pure liquids as in earlier work. There was no reaction of the solvent with oxidant. The rate constants, k<sub>obs</sub>, increases with decrease in dielectric constant of the medium. The plots of log k<sub>obs</sub> versus 1/D was linear (Fig. 1).
Effect of initially added products
The initially added products such as manganate, aldehyde and ammonia did not show any significant effect on the rate of the reaction.

Test for free radicals
The reaction mixture was kept for an hour in the acrylonitrile scavenger in an inert atmosphere. Diluting by methanol, the formation of precipitate indicates the presence of free radical in the reaction.

Effect of temperature
The rate of reaction was measured at different temperatures under varying [L(+)]-lysine. The rate constant, k of the slow step was obtained from the intercept of 1/k versus 1/[L(+)]-lysine and k was used to calculate the activation parameters. The values of k (s⁻¹) were 4.5 ± 0.2 x 10⁻³, 6.8 ± 0.3 x 10⁻⁴, 9.3 ± 0.4 x 10⁻⁴ and 12.8 ± 0.6 x 10⁻³ at 25, 30, 35 and 40°C respectively. The activation parameters corresponding to these constants were evaluated and are as follows: Eᵣ = 53 ± 2 kJ mol⁻¹, log A = 11.8 ± 0.5, ΔH* = 51 ± 2 kJ mol⁻¹, ΔG* = 60 ± 3 kJ mol⁻¹ and ΔS* = -28 ± 1 J K⁻¹ mol⁻¹.

The permanganate ion, MnO₄⁻, is a powerful oxidant in aqueous alkaline medium. As it exhibits multitude of oxidation states, the stoichiometric results and pH of the reaction media play a significant role. The Diode Array Rapid scan spectrophotometric (DARSS) studies have shown that at pH > 12 the product of the reaction of Mn(VII) is Mn(VI) and no further reduction was observed as reported by earlier workers. The formation of Mn(VI) was also evidenced by Fig. 2, where the absorption of Mn(VII) decreases at 526 nm and that of Mn(VI) increases at 608 nm during the course of reaction. However, on long standing Mn(VI) is slowly reduced to Mn(IV) under our experimental conditions.

The reaction between L(+)-lysine and permanganate in alkaline medium has a stoichiometry of 1:2 with a unit order in [permanganate] and less than unit order in both [L(+)-lysine] and [alkali]. No product effect was observed. In aqueous solution amino acids exist as zwitterionic form whereas in aqueous alkaline solution they exist as anionic form. The results suggest that first the alkali combines with permanganate to give an alkali-permanganate complex which then reacts with the substrate, L(+)-lysine, to give another complex. The latter complex decomposes in a slow step to give a free radical derived from decarboxylated L(+)-lysine, which further reacts with another molecule of permanganate in a fast step to yield the products (Scheme 3).

The evidence for complex formation was obtained from the UV-visible spectra of both MnO₄⁻ and MnO₄⁻ - L(+)lysine mixtures, in which a bathochromic shift of MnO₄⁻ from 318 to 345 nm and hyperchromicity at 345 nm was observed. This was also evidenced from the Michaelis Menton plot, and such complex formation between substrate and oxidant have also been observed in other studies. Since Scheme 3 is in accordance with the generally well accepted principle of non-complimentary oxidations taking place in sequences of one electron step, the reaction between the substrate and oxidant would afford a radical intermediate. A free radical scavenging experiment revealed such a possibility. This type of radical intermediate has also been observed in earlier work with amino acid.
Scheme 3 leads to rate law (2)

$$
\frac{d[\text{MnO}_4^-]}{dt} = \frac{kK_2[L\,\text{Lysine}][\text{OH}][\text{MnO}_4^-]}{1 + K_1[\text{OH}^+] + K_2[L\,\text{Lysine}][\text{OH}]} \quad (2)
$$

Strictly, the factors \((1 + K_1K_2[\text{MnO}_4^-][\text{OH}^+])\) and \((1 + K_1[\text{MnO}_4^-] + K_1K_2[\text{MnO}_4^-][\text{OH}])\) should also be in the denominator on RHS of Eq. (2) but in view of low concentration of Mn(VII) used and the observed first order in Mn(VII), these terms approximated to unity. Equation (2) can be rearranged to (3) which is suitable for verification.

$$
\frac{1}{k_{\text{on}}} = \frac{1}{kK_2[L\,\text{Lysine}][\text{OH}]} + \frac{1}{kK_2[L\,\text{Lysine}]} + \frac{1}{k} \quad (3)
$$

According to Eq. (3) the plots of \(1/k_{\text{on}}\) versus \(1/[L\,\text{Lysine}]\) and \(1/k_{\text{on}}\) versus \(1/[\text{OH}^-]\) should be linear and they are found to be so. The slopes and intercept of such plots lead to values of \(k\) at 25°C as \(4.5 \pm 0.2 \times 10^{-3}\) s\(^{-1}\), \(K_1 = 1.38 \pm 0.07 \times 10^3\) and \(K_2 = 2.1 \pm 0.1\) dm\(^3\) mol\(^{-1}\) respectively. Using these values, the rate constants over different experimental conditions were calculated and compared with experimental data in Table 1. There is a good agreement between them. The value of \(K_1\) is in agreement with earlier work\(^{23}\).

The effect of increase in ionic strength on the rate qualitatively explains the reaction between two negatively charged ions as seen in Scheme 3. However, increasing the content of \(t\)-butyl alcohol in the reaction medium leads to increase in the rate of reaction, contrary to the expected slower reaction between like ions in the media of lower relative permittivity. Perhaps this effect is countered substantially by the formation of active reaction species to a greater extent in a low relative permittivity media leading to the net increase in reaction rate\(^{23}\). The values of \(\Delta H^o\) and \(\Delta S^o\) were both favourable for electron transfer processes. Negative value of \(\Delta S^o\) within the range for radical reactions have been ascribed\(^{24}\) to the nature of electron pairing and electron unpairing processes and to the loss of degree of freedom, formerly available to the reactions on the formation of rigid transition state. The observed modest enthalpy of activation and relatively low value of entropy of activation and higher rate constant indicate that the oxidation presumably occurs by an inner sphere mechanism. This conclusion is supported by earlier work\(^{25}\).

It is also interesting to note that the oxidant species \([\text{MnO}_4^-]\) required a \(pH > 12\) and below which system gets disturbed and the reaction will proceed further to give reduced product of oxidant as Mn(IV) which develops tobacco coloured turbidity slowly. Thus it becomes apparent that in carrying out this reaction the role of \(pH\) in the reaction medium is crucial. It is also noteworthy that under the conditions studied, the reaction occurs in two successive one-electron reductions (Scheme 3) rather than two-electron reduction in a single step (Scheme 2).

References


