Chapter VI

OXIDATION OF TYROSINE BY HEXACYANOFERRATE(III) IN AQUEOUS ETHANOIC ACID MEDIUM - A KINETIC AND MECHANISTIC STUDY
Studies involving hexacyanoferrate(III) as an oxidant in acid media\textsuperscript{1-4} are limited by the fact that its reduction potential is small, 0.356 V, in such media\textsuperscript{5}. However, hexacyanoferrate(III) is very sensitive to the solvent and acidity. Increasing amount of ethanoic acid and hydrochloric acid in the medium lead to an increase in redox potential, the value reaches 1.14 V in 40\% ethanoic acid containing 5.0 mol dm\textsuperscript{-3} hydrochloric acid\textsuperscript{1,6} (Table VI (i) (p.192)). The difference of redox potential of the reactant in such media is reason enough to expect drastic changes in the identity of the reacting species in such solutions. Hexacyanoferrate(III) forms different species in acid media\textsuperscript{1,7}. Thus, by changing the solvent polarity and at high chloride concentration, the non-occurring reaction between tyrosine and hexacyanoferrate(III) becomes possible. The mechanism may be quite complicated due to the presence of high chloride and ethanoic acid content, and also due to the formation of different intermediates of oxidant. Hence we have investigated the kinetics of oxidation of tyrosine by hexacyanoferrate(III) in the presence of high concentration of hydrochloric and ethanoic acids in this chapter in order to understand the behavior of the active species of hexacyanoferrate(III) in such media and to arrive at a suitable mechanism.

**EXPERIMENTAL**

Reagent grade chemicals and double distilled water were used throughout the work. The stock solution of the oxidant, hexacyanoferrate(III), was prepared by dissolving potassium hexacyanoferrate(III) (BDH) in water and the concentration was ascertained by iodometric titration\textsuperscript{8}. A stock solution of tyrosine was prepared by dissolving tyrosine (s.d.fine chem.) in 1.0 mol dm\textsuperscript{-3} hydrochloric acid. Ethanoic acid and hydrochloric acid were obtained
Table VI (i)

Effect of hydrochloric and ethanoic acids on the reduction potentials of Fe(CN)$_6^{3-}$/Fe(CN)$_6^{4-}$ at 25° C

<table>
<thead>
<tr>
<th>% of Ethanoic acid – water (v/v)</th>
<th>Potentials of Fe(CN)$_6^{3-}$/Fe(CN)$_6^{4-}$ (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>00</td>
<td>0.357$^a$</td>
</tr>
<tr>
<td></td>
<td>0.982$^b$</td>
</tr>
<tr>
<td>10</td>
<td>1.00$^b$</td>
</tr>
<tr>
<td>20</td>
<td>1.03$^b$</td>
</tr>
<tr>
<td>30</td>
<td>1.13$^b$</td>
</tr>
<tr>
<td>40</td>
<td>1.14$^b$</td>
</tr>
</tbody>
</table>

$^a$ = Data from Ref. 5; in dilute acid medium

$^b$ = Data from Ref. 6; in presence of [HCl] = 5.0 mol dm$^{-3}$
from Glaxo Excelar. The former was also purified by refluxing with potassium permanganate for 5 h and then distilled. Further distillation yield the fraction at 118 °C to be used. Perchloric acid (70% E.Merck) and hydrochloric acid were the source of H⁺ and Cl⁻ ion utilized to vary the acid and chloride concentrations respectively in the reaction media. The ionic strength was maintained using sodium perchlorate. The sodium perchlorate solution was obtained by dissolving the required quantity of sodium perchlorate (Koch-Light) in water. hexacyanoferrate(II) solution was obtained by dissolving potassium hexacyanoferrate(II) (BDH) in water and standardizing with cerium(IV) solution.

Kinetic studies

Kinetic measurements in 40% ethanoic acid - water (v/v) and 4.0 mol dm⁻³ hydrochloric acid were carried out at 25 ± 0.1 °C and at I = 4.1 mol dm⁻³. Reactions were initiated by mixing previously thermostatted solutions of hexacyanoferrate(III) and tyrosine which also contained the required amounts of hydrochloric acid, ethanoic acid and sodium perchlorate. The kinetics were followed under pseudo-first order conditions with tyrosine in excess. Progress of the reaction was followed by measuring the absorbance of hexacyanoferrate(III) in the reaction mixture at 420 nm in a 1 cm cell placed in the thermostatted compartment of a Varian Cary 50 Bio UV-Vis spectrometer. At this wave length, all other materials concerned have negligible absorbance. Application of Beer's law under the reaction conditions had been verified (Fig. VI (i) (p.195)) between 1.0 x 10⁻⁴ and 1.0 x 10⁻³ mol dm⁻³ of
hexacyanoferrate(III) and e was found to be 1485 ± 25 dm³ mol⁻¹ cm⁻¹. The first order rate constants, $k_{obs}$ were obtained from the plot of log [hexacyanoferrate(III)] versus time. The first order plots were linear over three half-lives. The first order rate constants were average of minimum four sets and were reproducible within ±5%.

RESULTS

Stoichiometry and product analysis

Five different sets of reactant concentrations were mixed at constant concentration of acid and chloride (4.0 mol dm⁻³ in each case) and at ionic strength (4.1 mol dm⁻³). After 6 h, the oxidant, hexacyanoferrate(III) was analysed by measuring its absorbance at 420 nm. The product hexacyanoferrate(II) was determined with cerium(IV) solution. The stoichiometry was found to be two moles of hexacyanoferrate(III) reacted with one mole of tyrosine as shown in eqn. (1).

$$2\text{Fe(CN)}_6^{3-} + \text{H}_2\text{O} \rightarrow 2\text{Fe(CN)}_6^{4-} + \text{NH}_3 + \text{CO}_2 + 2\text{H}^+$$  

(1)

LC-ESI-MS analysis was carried out using reverse phase high performance liquid chromatography (HPLC) system with a phenomenes C-18 column, UV/Visible detector and series mass analyzer. 12 µL of acidified reaction mixture was injected. The mobile phase consisted of acetonitrile (eluent A) and methanol (Containing 0.1% CH₃COOH) at a flow rate of 1 ml/min. Gradient elution was run to separate substrate and reaction products. Gradient 0 min/ 95% A - 15 min/ 35% A - 25 min/ 35% A - 30 min/ 95% A -
Fig. VI (i)

Verification of Beer's law for hexacyanoferrate(III) concentrations at 420 nm in 4.0 mol dm$^{-3}$ hydrochloric and 40% ethanoic acids at 25 °C

![Graph showing the relationship between optical density and [Fe(CN)$_6^{3-}$] x 10$^4$ mol dm$^{-3}$]
35 min/ 35% A. LC-ESI-MS analysis of tyrosine reaction indicated the presence of two products with molecular ions of m/z 136 (Yield 85%) and 375 (Yield 05%) respectively. The molecular ion of tyrosine is m/z 182. The m/z 375 corresponds to dityrosine product and m/z 136 corresponds to p-hydroxy phenylacetaldehyde.

**Reaction order**

The orders with respect to each of oxidant, reductant and acid were determined by the slopes of log $k_{obs}$ versus log (concentration) plots. The orders were obtained by varying the concentrations oxidant, reductant and acid in turn while keeping all other concentrations and conditions constant.

**Effect of [hexacyanoferrate(III)]**

At constant concentration of tyrosine, $5.0 \times 10^{-3}$ mol dm$^{-3}$, hydrochloric acid, $4.0$ mol dm$^{-3}$, 40% ethanoic acid, and at constant ionic strength, $4.1$ mol dm$^{-3}$, the hexacyanoferrate(III) concentration was varied. The order with respect to hexacyanoferrate(III) in the range of $5.0 \times 10^{-5}$ to $5.0 \times 10^{-4}$ mol dm$^{-3}$ concentration range was unity. This was confirmed by the pseudo-first order rate constants which are constant at different hexacyanoferrate(III) concentrations (Table VI (ii) (p.198)). The pseudo-first order plots under these conditions were linear and parallel over 75% completion of the reaction at different initial concentrations of hexacyanoferrate(III) also confirming the unit order with respect to hexacyanoferrate(III) concentration.

**Effect of [tyrosine]**

The effect of tyrosine concentration on the reaction was studied at constant concentrations of acid, hexacyanoferrate(III), 40% ethanoic acid and at a
constant ionic strength of 4.1 mol dm$^{-3}$ at 25° C. The substrate, tyrosine was varied in the range of $1.0 \times 10^{-3}$ to $1.0 \times 10^{-2}$ mol dm$^{-3}$. The $k_{\text{obs}}$ values increased with increase in concentration of tyrosine (Table VI (ii) (p.198)). From the plot of log $k_{\text{obs}}$ versus log [tyrosine] the value of slope was found to be 0.71, indicating less than unit order with respect to tyrosine concentrations (Fig. VI (ii) (p.199)) ($r \geq 0.9959$, $S \leq 0.0023$).

**Effect of added product**

The effect of initially added product, hexacyanoferrate(II) was studied in the $5.0 \times 10^{-5} - 5.0 \times 10^{-4}$ mol dm$^{-3}$ concentration range, keeping the ionic strength, reactant concentrations and other conditions constant. It was found that added hexacyanoferrate(II) inhibits the rate of reaction (Table VI (ii) (p.198)). The order with respect to hexacyanoferrate(II) concentrations was found to be -0.95 (Fig. VI (ii) (199)). The inhibitory effect of initially added product, hexacyanoferrate(II) is shown in the Fig. VI (iii) (200) ($r \geq 0.992$, $S \leq 0.0031$).

**Effect of [acid]**

Increasing the acid concentration at a constant chloride concentration and constant ionic strength accelerates the reaction and the order with respect to $H^+$ ion concentration was determined from log - log plots of $k_{\text{obs}}$ versus concentrations and was found to be 1.96 in the 1.5 - 4.0 mol dm$^{-3}$ range of $H^+$ ion (Table VI (iii) (p.202)). A plot of $k_{\text{obs}} / [H^+]$ versus $[H^+]$ is linear (Fig. VI (iv) (p.203)) with a very small intercept, the slope is being nearly 10 times greater than the intercept. Thus, the reaction involving two protons, either as
Table VI (ii)

Effect of variation of hexacyanoferrate(III), tyrosine and hexacyanoferrate(II) concentrations on the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid - water (v/v) at 25 °C

\([\text{HCl}] = 4.0; \quad I = 4.10 / \text{mol dm}^{-3}\)

<table>
<thead>
<tr>
<th>([\text{Fe(CN)}_6^{3-}]\times10^4 ) (mol dm(^{-3}))</th>
<th>([\text{Tyrosine}] \times10^3 ) (mol dm(^{-3}))</th>
<th>([\text{Fe(CN)}_6^{4+}] \times10^4 ) (mol dm(^{-3}))</th>
<th>(k_{\text{obs}} \times10^4 ) (s(^{-1}))</th>
</tr>
</thead>
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</tr>
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<td>5.0</td>
<td>0.50</td>
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<td>3.0</td>
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<tr>
<td>2.0</td>
<td>5.0</td>
<td>5.0</td>
<td>0.97</td>
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</tbody>
</table>
Fig. VI (ii)

Order with respect to tyrosine and hexacyanoferrate(II) concentrations on the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid - water (v/v) at 25°C

(Conditions as in Table VI (ii) (p.198))

\[ 5 + \log [\text{Fe(CN)}_6^{4-}] \]

\[ 3 + \log [\text{Tyrosine}] \]
Inhibitory effect of hexacyanoferrate(II) concentrations on the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid - water (v/v) at 25°C.

(Conditions as in Table VI (ii) (p.198))
the monoprotonated oxidant species reacting with another free proton, or the
diprotonated oxidant species is likely to be the dominant path. Hexacyanoferrate(III) forms protonated species as shown in equilibria (2) to (4) with the stability constants \( \beta_1(= K_1) \), \( \beta_2(= K_1K_2) \) and \( \beta_3(= K_1K_2K_3) \) belonging to the species with one, two and three protons respectively\(^{1,9}\).

\[
\begin{align*}
\text{Fe(CN)}_6^{3-} + H^+ & \rightleftharpoons \text{HFe(CN)}_6^{2-} \quad K_1 \\
\text{HFe(CN)}_6^{2-} + H^+ & \rightleftharpoons \text{H}_2\text{Fe(CN)}_6^- \quad K_2 \\
\text{H}_2\text{Fe(CN)}_6^- + H^+ & \rightleftharpoons \text{H}_3\text{Fe(CN)}_6^3+ \quad K_3
\end{align*}
\]

With \([\text{Fe(CN)}_6^{3-}]_t\) and \([\text{Fe(CN)}_6^{3-}]_f\) representing the total and unprotonated hexacyanoferrate(III) species, the different oxidant species may be obtained from equation (5).

\[
[\text{Fe(CN)}_6^{3-}]_t = [\text{Fe(CN)}_6^{3-}]_f \left\{ 1 + \beta_1[H^+] + \beta_2[H^+]^2 + \beta_3[H^+]^3 \right\} \quad (5)
\]

The reported equilibrium constants of the different species are \( K_3 < K_2 < K_1 < 10 \), the value of \( K_1 \) being nearly 10 in the media under the present study\(^{1,9}\). While the singly protonated species is expected to form extensively in the high acid media used, the doubly and triply protonated species expected to form to a lesser extent.

The data of Table VI (iii) (p.202) are also used to obtain Fig. VI (v) (p.204), showing the plots of \( \log k_{\text{obs}} \) and also \( \alpha \), the fraction of \( \text{H}_2\text{Fe(CN)}_6^- \) of total \( \text{Fe(CN)}_6^{3-} \), as functions of \( \log [H^+] \). The absorbance of hexacyanoferrate(III) increases with increase in hydrogen ion concentration.
Table VI (iii)

Effect of variation of H⁺ ion concentration on the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid – water (v/v) at 25 °C

\[
\begin{align*}
[\text{Fe(CN)}_6^{3-}] &= 2.0 \times 10^{-4}; \\
[\text{Tyrosine}] &= 5.0 \times 10^{-3}; \\
[\text{HCl}] &= 4.0; \\
I &= 4.10 \text{ / mol dm}^{-3}
\end{align*}
\]

<table>
<thead>
<tr>
<th>[H⁺] (mol dm⁻³)</th>
<th>[Fe(CN)₆³⁻] × 10⁴</th>
<th>[HFe(CN)₆²⁻] × 10⁴</th>
<th>[H₂Fe(CN)₆⁺] × 10⁴</th>
<th>(\alpha)</th>
<th>(k_{\text{obs}}) × 10⁴ (s⁻¹)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.5</td>
<td>2.4</td>
<td>3.6</td>
<td>1.2</td>
<td>0.60</td>
<td>1.10</td>
</tr>
<tr>
<td>2.0</td>
<td>1.7</td>
<td>3.4</td>
<td>1.5</td>
<td>0.75</td>
<td>1.90</td>
</tr>
<tr>
<td>2.5</td>
<td>1.3</td>
<td>3.2</td>
<td>1.7</td>
<td>0.85</td>
<td>2.95</td>
</tr>
<tr>
<td>3.0</td>
<td>0.99</td>
<td>2.9</td>
<td>1.9</td>
<td>0.95</td>
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<tr>
<td>3.5</td>
<td>0.80</td>
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<td>2.1</td>
<td>1.05</td>
<td>5.89</td>
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<td>4.0</td>
<td>0.70</td>
<td>2.7</td>
<td>2.2</td>
<td>1.10</td>
<td>7.77</td>
</tr>
</tbody>
</table>

'\(\alpha\)' is the fraction of the total [Fe(CN)₆³⁻] of the H₂Fe(CN)₆⁺ species
Order with respect to H\(^+\) ion concentration on the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid - water (v/v) at 25 °C and a plot of \(k_{\text{obs}}/\text{[H}^+]\) versus \([\text{H}^+]\)

(Conditions as in Table VI (iii) (p.202))
Effect of acid concentration on different Fe(CN)$_6^{3-}$ species and also on the rate constants of the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid - water (v/v) at 25 °C. The concentrations of Fe(CN)$_6^{3-}$ and HFe(CN)$_6^{2-}$ decreases with increase of [H$^+$] and do not parallel the trend of reaction rate and hence are not shown.

(Conditions as in Table VI (iii) (p.202))
and finally remains constant. This indicates the possibility of various species of hexacyanoferrate(III) with hydrogen ion concentration, with predominance of one species of hexacyanoferrate(III) as H$_2$Fe(CN)$_6^+$. 

**Effect of [chloride]**

The effect of added chloride ions on the reaction in the concentration range, 0.25 to 4.0 mol dm$^{-3}$ was studied while keeping all the other concentrations constant. It was observed that added chloride did not have any significant effect on the pseudo-first order rate constants.

The experimental rate law is given by

\[
\text{Rate} = -\frac{d[\text{Fe(CN)}_6^{3-}]}{dt} = k [\text{Fe(CN)}_6^{3-}] [\text{tyrosine}]^{0.72} [\text{H}^+]^{1.96} [\text{Fe(CN)}_6^{4-}]^{-1} \quad (6)
\]

**Effect of dielectric constant and ionic strength**

At constant concentrations of reactants and with other conditions constant, the ionic strength was varied and found that ionic strength did not has any significant effect on the rate of reaction. The effect of dielectric constant (D) on the rate of reaction has been studied. As the ethanoic acid content of the medium increases from 0 - 60% (v/v), the rate of the reaction increases (Table VI (iv) (p.206). Since the dielectric constants of aqueous acetic acid are not available in the literature, they were computed from the pure liquid values as in the Chapter II (p.49). The plot of log $k_{obs}$ versus 1/D was linear with positive slope (Fig. VI (vi) (p.207)).
Table VI (iv)
Effect of variation of dielectric constant (D) on the oxidation of tyrosine by hexacyanoferrate(III) at 25 °C

\[
[\text{Fe(CN)}_6^{3-}] = 2.0 \times 10^{-4} \; ; \quad [\text{Tyrosine}] = 5.0 \times 10^{-3} \; ; \\
[\text{HCl}] = 4.0 \; ; \quad I = 4.10 / \text{mol dm}^{-3}
\]

<table>
<thead>
<tr>
<th>% of ethanoic acid - water (v/v)</th>
<th>D</th>
<th>(1/D \times 10^3)</th>
<th>(k_{\text{obs}} \times 10^4)</th>
</tr>
</thead>
<tbody>
<tr>
<td>00</td>
<td>78.50</td>
<td>12.74</td>
<td>0.86</td>
</tr>
<tr>
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</tr>
<tr>
<td>60</td>
<td>35.09</td>
<td>28.49</td>
<td>19.0</td>
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</tbody>
</table>
Fig. VI (vi)

Effect of variation of dielectric constant (D) on the oxidation of tyrosine by hexacyanoferrate(III) at 25 °C

(Conditions as in Table VI (iv) (p.206))
Test for free radicals (Polymerization study)

The intervention of free radicals was examined as follows: The reaction mixture, to which a known quantity of acrylonitrile scavenger had been added initially, was kept in an inert atmosphere for 1h at room temperature. When the reaction mixture was diluted with methanol no precipitate resulted, suggesting that there is no participation of free radicals in the reaction.

Effect of temperature

The influence of temperature on the reaction was studied at 25, 30, 35 and 40 °C by varying tyrosine and H⁺ ion concentrations. The rate constants were found to increase with increase in temperature. The rate constant k of the slow step of Scheme 1 were obtained from the slopes and intercepts of 1/k_{obs} versus 1/[tyrosine] and 1/k_{obs} versus 1/[H^+]² plots at four different temperatures. The data were subjected to least square analysis as in Chapter II (p.52) and are tabulated in Table VI (v) (p.209). The activation parameters corresponding to these rate constants were evaluated from the Arrhenius plot of log k versus 1/T (Fig. VI (vii) (p.210)) (r ≥0.999, S ≤ 0.001) for Scheme 1 and are given in Table VI (v) (p.209).

DISCUSSION

The effect of hydrogen ions on the reaction rate is to accelerate the reaction at constant chloride concentrations. The order with respect to acid is ca. 2 and suggests the involvement of one or more protonated species¹⁰. As mentioned in the result section (p.197) (effect of acid), since the order in acid concentration is very nearly two, the active form of oxidant is found to be as in
Table VI (v)

Activation and thermodynamic parameters of the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid - water (v/v)

(a) Effect of temperature

<table>
<thead>
<tr>
<th>Temperature (K)</th>
<th>k x 10^3 (s^-1)</th>
<th>log k (Y)</th>
<th>1/T x 10^3</th>
<th>log k Y* calc.</th>
</tr>
</thead>
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<td>-2.6801</td>
<td>298</td>
<td>-2.6823</td>
</tr>
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<td>-2.6151</td>
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(b) Activation Parameters

<table>
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<th>Parameters</th>
<th>Values</th>
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<td>E_a (kJ mol^-1)</td>
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<tr>
<td>ΔH^# (kJ mol^-1)</td>
<td>23.8 ± 0.2</td>
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<tr>
<td>ΔS^# (J K^-1 mol^-1)</td>
<td>-216 ± 6</td>
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<td>ΔG^# (kJ mol^-1)</td>
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(c) Equilibrium constants β and K_4 at different temperatures

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<th>K_4 x 10^2</th>
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(d) Thermodynamic quantities using β and K_4 values

<table>
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<th>Using K_4 values</th>
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<tbody>
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<td>ΔH (kJ mol^-1)</td>
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<td>-45.4 ± 1.5</td>
</tr>
<tr>
<td>ΔS (J K^-1 mol^-1)</td>
<td>252 ± 4</td>
<td>-110 ± 3</td>
</tr>
<tr>
<td>ΔG (kJ mol^-1)</td>
<td>6.8 ± 0.2</td>
<td>-11.7 ± 0.2</td>
</tr>
</tbody>
</table>
Fig. VI (vii)

Effect of variation of temperature on the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid - water (v/v)
The variation of rate with acidity was shown to parallel the trend of variation of concentration of the $H_2Fe(CN)_6^-$ species with acidity (Table VI (iii) (p.202); Fig. VI (iv) (p.203)). Hence the diprotonated species of the oxidant is considered to be the active species in high acid solutions. In most of the oxidation reactions, hexacyanoferrate(III) resembles copper(II), which involves free radical formation and rapidly oxidizes it.

Hexacyanoferrate(III) - hexacyanoferrate(II) system, which has higher redox potential than copper(II) - copper(I) system, substantiates a better possibility for the rapid oxidation of the free radical with hexacyanoferrate(III) and the rapid oxidation of free radicals might completely mask the polymerization.

\[
Fe(CN)_6^{3-} + 2[H^+] \overset{\beta}{\longrightarrow} H_2Fe(CN)_6^-
\]

\[
H_2Fe(CN)_6^- + \text{Complex (C)} \overset{K_4}{\rightarrow} \text{Complex (C)} \rightarrow \text{slow} \rightarrow \text{fast} \rightarrow
\]

\[
\begin{align*}
\text{H}_2\text{Fe(CN)}_6^- & + \text{H}_2\text{O} \rightarrow \text{NH}_3 + \text{H}_2\text{Fe(CN)}_6^{2-} + \text{H}^+ \\
\text{Ammonia} & + \text{Hexacyanoferrate(II)} + \text{Water} \rightarrow \text{Ammonium} + \text{Hexacyanoferrate(II)} + \text{Proton}
\end{align*}
\]

**Scheme 1**

From scheme 1 the rate law (11) can be derived as follows:

\[
\text{Rate} = - \frac{d[Fe(CN)_6^{3-}]}{dt} = k \beta K_4 \text{[Tyrosine]} r [Fe(CN)_6^{3-}]_r [H^+]^2_r 
\]

(7)

The total $[Fe(CN)_6^{3-}]$ can be written as
\[ \text{[Fe(CN)}_6^{3-}\text{]}_t = [\text{Fe(CN)}_6^{3-}\text{]}_f + [\text{H}_2\text{Fe(CN)}_6^-] + \text{[Complex (c)]} \]
\[ = [\text{Fe(CN)}_6^{3-}\text{]}_f + \beta [\text{Fe(CN)}_6^{3-}] [\text{H}^+]^2 + \beta K_4 [\text{Tyrosine}] [\text{Fe(CN)}_6^{3-}] [\text{H}^+]^2 \]
\[ = [\text{Fe(CN)}_6^{3-}\text{]}_f \{1 + \beta [\text{H}^+]^2 + \beta K_4 [\text{Tyrosine}] [\text{H}^+]^2\} \]

Where 't' and 'f' stands for total and free

\[ [\text{Fe(CN)}_6^{3-}\text{]}_f = \frac{[\text{Fe(CN)}_6^{3-}\text{]}_t}{1 + \beta [\text{H}^+]^2 + \beta K_4 [\text{Tyrosine}] [\text{H}^+]^2} \] (8)

Similarly, total [Tyrosine] can be calculated as,

\[ [\text{Tyrosine}]_t = [\text{Tyrosine}]_f + \text{[Complex (c)]} \]
\[ = [\text{Tyrosine}]_f + \beta K_4 [\text{Tyrosine}] [\text{Fe(CN)}_6^{3-}] [\text{H}^+]^2 \]
\[ = [\text{Tyrosine}]_f \{1 + \beta K_4 [\text{Fe(CN)}_6^{3-}] [\text{H}^+]^2\} \]
\[ [\text{Tyrosine}]_f = \frac{[\text{Tyrosine}]_t}{1 + \beta K_4 [\text{Fe(CN)}_6^{3-}] [\text{H}^+]^2} \]

In view of low concentration of [Fe(CN)_6^{3-}] used, above equation can be written as

\[ [\text{Tyrosine}]_t = [\text{Tyrosine}]_f \] (9)

Similarly,

\[ [\text{H}^+]_t = [\text{H}^+]_f \] (10)

Substituting equations (8), (9) and (10) in equation (7) we get,

\[ \text{Rate} = -\frac{d [\text{Fe(CN)}_6^{3-}]}{dt} = \frac{k \beta K_4 [\text{Tyrosine}] [\text{Fe(CN)}_6^{3-}] [\text{H}^+]^2}{1 + \beta [\text{H}^+]^2 + \beta K_4 [\text{Tyrosine}] [\text{H}^+]^2} \]

or

\[ \frac{\text{Rate}}{[\text{Fe(CN)}_6^{3-}]} = k_{\text{obs}} = \frac{k \beta K_4 [\text{Tyrosine}] [\text{H}^+]^2}{1 + \beta [\text{H}^+]^2 + \beta K_4 [\text{Tyrosine}] [\text{H}^+]^2} \] (11)
It has been observed that the presence of ferrocyanide, \( \text{Fe(CN)}_6^{4-} \) retards the rate of reduction of ferricyanide by tyrosine. However, the retardation is small, and this has been attributed to a secondary salt effect\(^1\). Because of high acidity, catalytically more active protonated species of ferrocyanide ions produced in the system cause the secondary salt effect. Hence in the rate law (11), the term \([\text{Fe(CN)}_6^{4-}]\) is not included. For verification, rate law (11) is rearranged to (12)

\[
\frac{1}{k_{\text{obs}}} = \frac{1}{k \beta K_4 [\text{Tyrosine}] [\text{H}^+]^2} + \frac{1}{k K_4 [\text{Tyrosine}]} + \frac{1}{k} \quad (12)
\]

According to Eq. (12), other conditions being constant, plots of \(1/k_{\text{obs}}\) versus \(1/[\text{Tyrosine}]\) \((r \geq 0.9902, S \leq 0.015)\) and \(1/k_{\text{obs}}\) versus \(1/[\text{H}^+]^2\) \((r \geq 0.9988, S \leq 0.2)\) be linear and are found to be so (Fig. VI (viii) (p.214)). The slopes and intercepts of such plots lead to the values of \(\beta, K_4\) and \(k\) (Table VI (v) (p.209)).

The thermodynamic quantities for the first and second equilibrium steps of Scheme 1 can be evaluated as follows. The tyrosine and \(\text{H}^+\) ion were varied at four different temperatures. The plots of \(1/k_{\text{obs}}\) versus \(1/\text{Tyrosine}\) and \(1/k_{\text{obs}}\) versus \(1/[\text{H}^+]^2\) should be linear. From the slopes and intercepts, the values of \(\beta\) and \(K_4\) were calculated at different temperatures and these values are given in Table VI (v) (p.209). van’t Hoff plots were made for variation of \(\beta\) and \(K_4\) with temperature \((\log \beta \text{ versus } 1/T \ (r \geq 0.9889, S \leq 0.007)\) and \(\log K_4 \text{ versus } 1/T \ (r \geq 0.9905, S \leq 0.0023)\). The values of enthalpy of reaction, \(\Delta H\), entropy of reaction, \(\Delta S\) and free energy of reaction, \(\Delta G\), were calculated for the first and second equilibrium steps. These values are given in Table VI (v) (p.209).
Fig. VI (viii)

Verification of rate law (11) in the form of (12) on the oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid - water (v/v) at 25 °C

(Conditions as in Table VI (ii) (p.198) and Table VI (iii) (p.202))
The moderate values of $\Delta H^\#$ and $\Delta S^\#$ were both favorable for electron transfer processes. The negative value of $\Delta S^\#$ indicates that the complex(C) is more ordered than the reactants$^{14}$.

The effect of ionic strength on the rate qualitatively explains the reaction between negatively charged ion and neutral molecule. The effect of solvent on the reaction rate has been described in detail in the literature$^{1,15}$. The increase in the ethanoic acid content in the reaction medium leads to an increase in the rate of reaction, contrary to the expected between negatively charged and the neutral molecule in the media of lower dielectric constant. Perhaps this effect is countered substantially by the formation of active reactant species to greater extent in low dielectric constant, leading to net increase in reaction rate$^{16}$. This also supports the enhancement of reduction potentials of Fe(CN)$_6^{3-}$/Fe(CN)$_6^{4-}$.

**CONCLUSION**

It is apparent that the reaction between tyrosine and hexacyanoferrate(III) in 40% ethanoic acid - water (v/v) containing large quantities of hydrochloric acid (4.0 mol dm$^{-3}$) is rendered possible by the difference in reduction potentials brought about due to the facile formation of species of hexacyanoferrate(III). $\text{H}_2\text{Fe(CN)}_6^-$ is the main active species of the reactant, although the other species might be active to much lesser extent. The dielectric constant of the medium plays crucial role in the reaction.

**FINDINGS OF THE CHAPTER VI**

The kinetics of oxidation of tyrosine by hexacyanoferrate(III) in 40% ethanoic acid – water containing 4.0 mol dm$^{-3}$ hydrochloric acid exhibited 2:1
stochiometry. The oxidation is made possible by varying the difference in redox potential by adding ethanoic acid to the reaction mixture. The reaction is first order with respect to hexacyanoferrate(III) and fractional (0.72) order with respect to tyrosine concentrations. Increasing the acid concentration at constant chloride concentration accelerates the reaction. The order with respect to acid concentration is nearly two. Amongst the initially added products hexacyanoferrate(II) retards the reaction. $\text{H}_2\text{Fe(CN)}_6^-$ is the active species of hexacyanoferrate(III). A suitable mechanism is proposed. The activation parameters and thermodynamic quantities with respect to slow step of Scheme 1 were calculated and discussed.
REFERENCES

1. S. A. Chimatadar, M.S. Salunke and S. T. Nandibewoor,
   Transition Met. Chem., 29, 743 (2004);
   S. A. Chimatadar, S. B. Koujalagi and S. T. Nandibewoor,

2. V. K. Jindal and M. C. Agrawal,

3. L. Adamcikova and L. Treindl,

4. C. W. J. Scaif and R. G. Wilkins,

5. M. C. Day, Jr. and J. Selbin,
   Theoretical Inorganic Chemistry, Reinhold Publishing Corporation,
   New York, (1964) p.226

6. N. Rukmini, V. S. N. P. Kavitha and K. V. Rama Rao,
   Talanta, 26, 579 (1979)

7. S. A. Chimatadar, S. T. Nandibewoor and J. R. Raju,

8. G. H. Jeffery, J. Bassett, J. Mendham and R. C. Denney,
   Vogel's Text Book of Quantitative Chemical Analysis, ELBS,

9. J. Jordan and G. J. Ewing,
   Inorg. Chem., 1, 587 (1962);
   L. D. Pedro, G. Begona and M. L. Jose,

10. K. S. Gupta and Y. K. Gupta,
11. M.P. Singh and S. Gosh,
Z. Phys. Chem. (Leipzig), 204, 1 (1955);
M.P. Singh and S. Gosh,
Z. Phys. Chem. (Leipzig), 207, 187 (1957);
S.A. Farokhi and S. T. Nandibewoor

12. K. B. Wiberg and W. G. Nigh,

13. J. K. Kochi, B.M. Graybill and M. Kwiz,
J. Am. Chem. Soc., 86, 5257 (1964)

14. Weissberger, (Ed.),

15. K. J. Laidler,
E. S. Amis,

16. S. T. Nandibewoor and V. A. Morab,