Kinetics and Mechanism of Benzoin Oxidation Using Iron(III) in the Presence of Ferrozine or 2,2'-Bipyridine in Acidic Medium

KHALDOUN A. AL-SOU'OD, BASEM F. ALI, RAJAB ABU-EL-HALAWA, ABD-AL-HAKEEM H. ABU-NAWAS

Chemistry Department, Al al-Bayt University, Mafraq, Jordan

Received 6 December 2004; accepted 25 February 2005

DOI 10.1002/kin.20092

Published online in Wiley InterScience (www.interscience.wiley.com).

ABSTRACT: Kinetics and mechanism of oxidation of benzoin (H₂B) by ferrozine (Fz) or 2,2'bipyridine (bipy) have been carried out in aqueous HNO₃ medium. The rate shows first-order dependence on [H₂B] and [Fe³⁺] and inverse second-order dependence on [H⁺]. The rate of reaction increased with increase in the ligand concentration. The increase in dielectric constant will increase the rate, while increase in [HNO₃] decreased it. Substituent and temperature effects on the rates have been investigated. The rate laws derived are in excellent agreement with the experimental results. Plausible mechanisms are suggested. © 2005 Wiley Periodicals, Inc. Int J Chem Kinet 37: 444–449, 2005

INTRODUCTION

The oxidation of benzoin (H₂B) to benzil by a variety of oxidizing agents such as 1,10-phenanthrolein [1], sulfur [2], bismuth triacetate [3], hexacyanoferrate(III) [4], and nitric acid [5] has been a subject of a considerable number of studies [1–10]. The mechanisms and kinetics of the oxidation of benzoins have been previously studied [11–14]. Oxidation of organic compounds by complex ion may occur via free radical and further oxidation to yield the products [15–17]. Also, the oxidation processes may involve electron

© 2005 Wiley Periodicals, Inc.

transfer through inner sphere [18,19] or outer sphere [20] mechanisms. In a continuation of our work on benzoin, we report herein the oxidation kinetics of benzoin with iron(III) in the presence of ferrozine (Fz) or 2,2'-bipyridine (bipy) in acidic medium. The structure of ferrozine is shown in Fig. 1.

EXPERIMENTAL

Materials

Benzoin (H₂B) derivatives were prepared as reported in the literature [21–23]. All other chemicals used were of analytical grade. Solutions were prepared in doubly distilled deionized water (second distillation being from permanganate). All glass vessels were Pyrex.

Correspondence to: K. A. Al-Sou'od; e-mail: khaldoun@aabu. edu.jo.

Contract grant sponsor: Al al-Bayt University.



Kinetic Measurements

The kinetic measurements were carried out at 30°C in HNO₃ acidic medium under the conditions [Fe³⁺], [Fz], [bipy], [H₂B]. All runs were carried out under pseudo-first-order conditions of $[H_2B] \gg [Fe^{3+}]$ and at constant ionic strength (I) maintained by the addition of sodium nitrate. The progress of the reaction was monitored by measuring the absorbance of the formed complexes at 562 and 255 nm against time for iron(III)ferrozine and iron(III)-2,2'-bipyridine, using a Unicam UV/VIS spectrophotometer with 1.0 cm cell. The isolation method was used in this study where the concentration of the substance to be monitored is changed with the others kept constant. Aqueous solutions are prepared by making a constant total volume for all runs. These solutions were made by mixing appropriate amount of reagents in a sequence (Fe(III) first, followed by HNO₃ and then NaNO₃, ligand, and finally benzoin). The blank solution is prepared in a similar manner with no benzoin added. The solutions and blank were equilibrated at 30°C for 24 h (completion of reaction) and then absorbance is measured (A_{∞}) .

The pseudo-first-order rate constant (k_{obs}) was obtained from the slope of the $\ln(A_{\infty} - A_t)$ against time plot (where A_{∞} and A_t are absorbances at infinite time and at *t* respectively). The temperature was kept constant using a liquid circulating bath C85A.

RESULTS AND DISCUSSION

Stoichiometry and Product Analysis

Known amount of H_2B reacted completely with excess of iron(III) at 30°C in 0.01 M HNO₃ (50% aqueous ethanol (v/v)) in the presence of excess Fz or bipy for 24 h. Estimation of unreacted iron(III) showed that 2 moles of iron(III) was consumed for each mol of H_2B in accordance with the following equation:



The resulting benzil (R = H) from oxidation was separated by extraction using CHCl₃ and then by evaporation of solvent. The product recrystallized from EtOH/H₂O. For (C₆H₅)–C(O)C(O)–(C₆H₅), m.p. = 92–94°C (lit. m.p. = 94.8°C); UV/VIS (EtOH): λ_{max} (log ε) 256 nm (4.24). Qualitative identification of benzil was also made from TLC by comparing the $R_{\rm f}$ values for the product from the oxidation reaction and a standard pure sample of benzil.

Iron(III) Dependence

The kinetic measurements were carried out under pseudo-first-order conditions. Initial concentration of iron(III) was 2.0×10^{-4} M. A plot of $\ln(A_{\infty} - A_t)$ against time was linear indicating a first-order dependence on iron(III) (see Fig. 2).

H₂B Dependence

Isolation method was used to determine the order with respect to H₂B. Different initial concentrations of H₂B were prepared while keeping the other parameters constant. It was found that k_{obs} increased as [H₂B] increased (Table I). Plot log k_{obs} versus log[H₂B] was linear with a slope = 1, indicating a first order with respect to H₂B. Plot of $1/k_{obs}$ versus $1/[H_2B]$ was linear with a positive intercept (Fig. 3).

Ligands Dependence

Effect of ligand on the pseudo-first-order rate constant (k_{obs}) was studied at 30°C by varying the ligand concentrations from 3.0×10^{-3} to 7.0×10^{-3} M and keeping the other parameters constant. Rate constant increases with increase in ligand concentrations (Table I). These kinetic results reflect the higher



Figure 2 Plot of $\ln(A_{\infty} - A_t)$ vs. time at $[Fe^{3+}] = 2.0 \times 10^{-4}$ M, $[H_2B] = 2.0 \times 10^{-3}$ M, $[L] = 7.0 \times 10^{-3}$ M, $[H^+] = 1.0 \times 10^{-2}$ M, I = 0.06 M, 50% ethanol-water (v/v) and 30°C.

(v/v)					
$[{ m H}_2{ m B}] \times 10^3 \ ({ m M})$	[bipy] × 10 ³ (M)	$[{ m H^+}] \times 10^2 ({ m M})$	$k_{\rm obs}~({\rm min}^{-1})$		
2.0	7.0	1.0	$2.02 \times 10^{-3} \pm 5.59 \times 10^{-5}$		
3.0	7.0	1.0	$3.01 \times 10^{-3} \pm 5.77 \times 10^{-5}$		
4.0	7.0	1.0	$3.93 \times 10^{-3} \pm 6.35 \times 10^{-5}$		
5.0	7.0	1.0	$4.80 \times 10^{-3} \pm 8.67 \times 10^{-5}$		
6.0	7.0	1.0	$5.74 \times 10^{-3} \pm 8.93 \times 10^{-5}$		
4.0	3.0	1.0	$1.71 \times 10^{-3} \pm 4.88 \times 10^{-5}$		
4.0	4.0	1.0	$2.44 \times 10^{-3} \pm 6.63 \times 10^{-5}$		
4.0	5.0	1.0	$3.10 \times 10^{-3} \pm 9.01 \times 10^{-5}$		
4.0	6.0	1.0	$3.63 \times 10^{-3} \pm 4.81 \times 10^{-5}$		
4.0	7.0	1.0	$4.01 \times 10^{-3} \pm 7.08 \times 10^{-5}$		
4.0	4.0	1.0	$3.73 \times 10^{-3} \pm 3.36 \times 10^{-5}$		
4.0	4.0	1.4	$3.13 \times 10^{-3} \pm 5.71 \times 10^{-5}$		
4.0	4.0	2.0	$2.71 \times 10^{-3} \pm 6.34 \times 10^{-5}$		
4.0	4.0	2.6	$2.32 \times 10^{-3} \pm 6.98 \times 10^{-5}$		
4.0	4.0	3.2	$2.01 \times 10^{-3} \pm 6.01 \times 10^{-5}$		

Table I Effect of $[H_2B]$, [L], and $[H^+]$ on k_{obs} at 30°C. $[Fe^{3+}] = 2.0 \times 10^{-4}$ M, I = 0.06 M, 50% ethanol-water (v/v)

oxidizing ability of the 1:2 formed complex $[FeL_2]^{2+}$ compared to uncomplexed Fe(III). This may be due to the ability of ligand to increase the stability of the complex in the lower oxidation state of iron $[FeL_2]^{2+}$ over the higher one $[FeL_2]^{3+}$, through back donation [15]. The plots of $1/k_{obs}$ versus $1/[L]^2$ were linear with positive slopes and intercepts (Fig. 4).

Hydrogen Ion Dependence

The values of k_{obs} were determined at different hydrogen ion concentrations (Table I) with ionic strength and other reactants concentrations kept constant. k_{obs} was found to decrease with increase in acid concentration.



Figure 3 Plot of $1/k_{obs}$ vs. $1/[H_2B]$ at $[H^+] = 1.0 \times 10^{-2}$ M, $[Fe^{3+}] = 2.0 \times 10^{-4}$, I = 0.06 M, 50% ethanol-water (v/v) and 30°C.



Figure 4 Plot of $1/k_{obs}$ vs. $1/[L]^2$ at $[Fe^{3+}] = 2.0 \times 10^{-4}$, $[H^+] = 1.0 \times 10^{-2}$ M, I = 0.06 M, 50% ethanol-water (v/v) and 30°C.

The plot of $1/k_{obs}$ versus $[H^+]^2$ was linear with a positive intercept (Fig. 5).

Substituent Effects

Substituent effects on the oxidation rates of H₂B have been studied at 30°C (Table II). It was found that the electron-withdrawing groups enhance the oxidation rate, whereas the electron-donating groups inhibit it. A plot of $\log(k/k_{\rm H})$ versus substituent constant (σ) was linear (Fig. 6) in accord with Hammett's equation. The reaction constant (ρ) was equal to \approx 1.2. This indicates that the reaction site has a higher electron density in the transition state compared to the starting material.

Ionic Strength Dependence

The influence of ionic strength on k_{obs} was studied at 30°C. Different initial concentrations of KNO₃ were prepared while keeping the other parameters constant.



Figure 5 Plot of $1/k_{obs}$ vs. $[H^+]^2$ at $[H_2B] = 4.0 \times 10^{-3}$ M, I = 0.06 M, $[Fe^{3+}] = 2.0 \times 10^{-4}$, 50% ethanol-water (v/v) and 30°C.

(15//1)		()	
Substrate	Substituent Constant (σ)	$k_{\rm obs} \times 10^3 ({\rm min}^{-1}) ({\rm Fz})$	$k_{\rm obs} ({\rm min}^{-1}) ({\rm bipy})$
Н	0.0	$3.41 \times 10^{-3} \pm 6.64 \times 10^{-5}$	$4.02 \times 10^{-3} \pm 7.32 \times 10^{-5}$
<i>m,m</i> ′-CH ₃ O	0.11	$5.13 \times 10^{-3} \pm 6.99 \times 10^{-5}$	$5.31 \times 10^{-3} \pm 6.59 \times 10^{-5}$
<i>m,m</i> ′-Cl	0.37	$1.14 \times 10^{-3} \pm 4.72 \times 10^{-5}$	$1.20 \times 10^{-3} \pm 3.05 \times 10^{-5}$
<i>pp</i> ′-Cl	0.22	$7.20 \times 10^{-3} \pm 8.43 \times 10^{-5}$	$7.64 \times 10^{-3} \pm 9.11 \times 10^{-5}$
p, p'-CH ₃	-0.17	$2.41 \times 10^{-3} \pm 5.51 \times 10^{-5}$	$2.61 \times 10^{-3} \pm 4.44 \times 10^{-5}$
<i>p</i> , <i>p</i> ′-CH ₃ O	-0.28	$1.93 \times 10^{-3} \pm 5.66 \times 10^{-5}$	$2.13 \times 10^{-3} \pm 3.81 \times 10^{-5}$

Table II Substituent Effect on k_{obs} at $[Fe^{3+}] = 2.0 \times 10^{-4}$ M, $[H_2B] = 4.0 \times 10^{-3}$ M, [L] = 0.004 M (Fz), and 0.007 M (bipy), $[H^+] = 0.01$ M, I = 0.06 M, 50% ethanol-water (v/v) at 30°C

From data in Table III it can be observed that higher k_{obs} are attained at higher ionic strengths. Using k_{obs} data at different ionic strengths, a graph was obtained employing the Brønsted–Bjerrum equation [24] based on the extended law of Debye–Hückel (Fig. 7). The rate constant at infinite dilution (k_0) is 1.32×10^{-3} and 1.24×10^{-3} min⁻¹ for Fz and bipy, respectively. The product of the charges $Z_A Z_B \approx 3$. This is the same expected value considering the proposed mechanism.

Thermodynamic Parameters

The reaction was studied at different temperatures (298–313 K). Plots of $\ln k_{obs}$ and $\ln(k_{obs}.h/k_{\rm B}T)$ versus 1/T yield the activation energy E_a , enthalpy of activation ΔH^{\neq} , entropy of activation ΔS^{\neq} , and Gibbs energy of activation ΔG^{\neq} (Table IV).

Since the pK_a values for bipy [25] and Fz [26] are 4.33 and 3.13, respectively, they exist mainly as protonated species (HL⁺) in the examined acid range (0.01–0.018 M). It is reasonable to consider that the first step in the mechanism involves the fast formation of the ligand protonated species (Scheme 1). In the present study, the Fe(III) concentration employed was on the order of



Figure 6 Plot of σ vs. $\log(k_{obs}/k_{\rm H})$ at $[{\rm Fe}^{3+}] = 2.0 \times 10^{-4}$ M, $[{\rm H}_2{\rm B}] = 4.0 \times 10^{-3}$ M, $[{\rm L}] = 0.004$ M (Fz), and 0.007 M (bipy), $[{\rm H}^+] = 0.01$ M, ${\rm I} = 0.06$ M, 50% ethanol-water (v/v) at 30°C.

10–4 M while the hydrogen ion concentration was in the range 0.01–0.018 M. Therefore, it is quite reasonable to assume that there is no appreciable dimerization under these conditions [27] and the concentration of hydroxylated Fe(III) is neglected due to low-hydrolysis constant. The hydrolysis is also suppressed, since the ligand (Fz or bipy) was present in large excess. The second step in the mechanism may proceed by the outer sphere route (step 3a in Scheme 1) or by inner sphere reaction (step 3b in Scheme 1). Steps 4 and 5 describe the formation of final products. In step 4, the formed radicals HB⁻ react with 1 mol of $[FeL_2]^{3+}$ to form benzil (B). Then in step 5, 1 mol of $[FeL_2]^{2+}$ reacts with another mol of HL⁺ to form a more stable complex $[FeL_3]^{2+}$.

The derived rate laws from the proposed mechanism are given below.

For Inner Sphere Mechanism

Rate =
$$\frac{-d[Fe^{3+}]}{dt}$$

= $\frac{kK_1K_2[L]^2[Fe^{3+}][H_2B]}{K_1[L]^2 + K_1K_2[H_2B][L]^2 + [H^+]^2}$ (1)

1.
$$H^+ + L \xrightarrow{Fast} HL^+$$

2. $Fe^{3+} + 2HL^+ \xrightarrow{K_1} [FeL_2]^{3+} + 2H^+$
3a. $[FeL_2]^{3+} + H_2B \xrightarrow{k}_{slow} [FeL_2]^{2+} + H^+ + HB^-$
or

3b.
$$[\operatorname{FeL}_2]^{3+} + \operatorname{H}_2B \xrightarrow{k_2} [\operatorname{FeL}_2BH_2]^{3+}$$

 $[\operatorname{FeL}_2BH_2]^{3+} \xrightarrow{k} [\operatorname{FeL}_2] + \operatorname{H}^+ + \operatorname{HB}^-$
4. $\operatorname{HB} + [\operatorname{FeL}_2]^{3+} \xrightarrow{\operatorname{Fast}} [\operatorname{FeL}_2]^{2+} + \operatorname{H}^+ + \operatorname{B}$
5. $[\operatorname{FeL}_2]^{2+} + \operatorname{HL}^+ \xrightarrow{\operatorname{Fast}} [\operatorname{FeL}_3]^{2+} + \operatorname{H}^+$

Table III Effect of Ionic Strength on k_{obs} at $[Fe^{3+}] = 2.0 \times 10^{-4}$ M, $[H_2B] = 4.0 \times 10^{-3}$ M, [L] = 0.004 M (Fz), and 0.007 M (bipy), $[H^+] = 0.01$ M, I = 0.06 M, 50% ethanol–water (v/v) at 30°C

Ionic					
Strength					
(M)	0.020	0.030	0.040	0.050	
$\frac{k_{\rm obs}({\rm bipy})}{{\rm min}^{-1}}$	$2.60 \times 10^{-3} \pm 4.98 \times 10^{-5}$	$3.14 \times 10^{-3} \pm 5.79 \times 10^{-5}$	$3.44 \times 10^{-3} \pm 7.35 \times 10^{-5}$	$3.74 \times 10^{-3} \pm 6.01 \times 10^{-5}$	
$k_{\rm obs}({\rm Fz})$ min ⁻¹	$2.84 \times 10^{-3} \pm 3.18 \times 10^{-5}$	$3.39 \times 10^{-3} \pm 4.71 \times 10^{-5}$	$3.81 \times 10^{-3} \pm 5.99 \times 10^{-5}$	$4.1 \times 10^{-3} \pm 7.64 \times 10^{-5}$	

according to Eq. (1), k_{obs} can be obtained:

$$k_{\rm obs} = \frac{kK_1K_2[L]^2[H_2B]}{K_1[L]^2 + K_1K_2[H_2B][L]^2 + [H^+]^2}$$
(2)

the inverse of k_{obs} is given by Eq. (3):

$$\frac{1}{k_{\text{obs}}} = \frac{1}{k} + \frac{1}{kK_2[\text{H}_2\text{B}]} + \frac{[\text{H}^+]^2}{kK_1K_2[\text{L}]^2[\text{H}_2\text{B}]}$$
$$= \frac{1}{k} + \frac{1}{[\text{H}_2\text{B}]} \left[\frac{1}{kK_2} + \frac{[\text{H}^+]^2}{kK_1K_2[\text{L}]^2} \right] \quad (3)$$

For *outer sphere mechanism*, the rate law, k_{obs} and its inverse are given by Eqs. (4)–(6).

Rate =
$$\frac{kK_1[L]^2[Fe^{3+}][H_2B]}{K_1[L]^2 + [H^+]^2}$$
 (4)

$$k_{\rm obs} = \frac{kK_1[\rm L]^2[\rm H_2B]}{K_1[\rm L]^2 + [\rm H^+]^2}$$
(5)

$$\frac{1}{k_{\text{obs}}} = \frac{1}{k[\text{H}_2\text{B}]} + \frac{[\text{H}^+]^2}{kK_1[\text{L}]^2[\text{H}_2\text{B}]}$$
$$= \frac{1}{[\text{H}_2\text{B}]} \left[\frac{1}{k} + \frac{[\text{H}^+]^2}{kK_1[\text{L}]^2} \right]$$
(6)



Figure 7 Plot of ionic strength vs log k_{obs} at $[Fe^{3+}] = 2.0 \times 10^{-4} \text{ M}, [H_2B] = 4.0 \times 10^{-3} \text{ M}, [H^+] = 0.01 \text{ M}, I = 0.06 \text{ M}, 50\%$ ethanol-water (v/v) at 30°C.

This rate law demands that the plots of $1/k_{obs}$ versus $1/[H_2B]$, $1/k_{obs}$ versus $1/[L]^2$, and $1/k_{obs}$ versus $[H^+]^2$ should be all linear with positive intercepts on the $1/k_{obs}$ axis. This is actually found to be the case as shown in Figs. 3–5 suggesting the compatibility of inner sphere mechanism with experimental results (Eq. (3)). Values of k, K_1 , and K_2 calculated from the intercepts and slopes of lines in Figs. 3 and 5 are listed in Table IV.

The estimated activation parameters (Table IV) indicate that the oxidation rate of benzoin readily occurs in the presence of bipy compared to Fz. The large negative ΔS^{\neq} values indicate the formation of the compact activated complex in accord with inner sphere mechanism.

CONCLUSION

Oxidation of benzoin in acidic medium using Fe(III) in the presence of Fz and bipy was studied along with kinetics and mechanisms. Kinetics was carried out under pseudo-first-order conditions and at constant ionic strength. The order of oxidation reaction with respect to reactants was studied and found to be of a first order with respect to both Fe(III) and H₂B. Moreover, the effect of ligand as well as the hydrogen ion concentrations on the pseudo-first-order rate constant was examined. These studies show that the rate constant increases with increase in ligand concentration, while it decreases with increase in [H⁺]. The H₂B substituent effects on the oxidation were also investigated in accord with Hammett's equation. Results revealed that electron-withdrawing groups enhance oxidation, while the electron-donating groups inhibit it. Furthermore, the ionic strength dependence was studied using the Brønsted-Bjerrum equation. The higher the ionic strength the larger is the rate constant. According to the results deduced from plots of $1/k_{obs}$ versus $1/[H_2B]$, $1/k_{obs}$ versus $1/[L]^2$, and $1/k_{obs}$ versus $[H^+]^2$ (Figs. 3–5) and the estimated activation parameters (Table IV), the

Complex	<i>E_a</i> (J/mol K)	ΔH^{\neq} (kJ/mol)	ΔS^{\neq} (J/mol K)	ΔG^{\neq} (kJ/mol)	$k \pmod{(\min^{-1})}$	$\begin{array}{c} K_1 \\ (\mathrm{M}^{-1}) \end{array}$	K_2 (M ⁻¹)
$[Fe(Fz)_2]^{3+}$	55.31	52.77	-148.09	98.04	0.0448	8.514	40.966
$[Fe(bipy)_2]^{3+}$	49.70	47.16	-169.3	98.92	0.0923	9.869	16.827

Table IV Values of k, K_1 , K_2 , E_a , ΔH^{\neq} , ΔS^{\neq} and ΔG^{\neq} for $[Fe(Fz)_2]^{3+}$ and $[Fe(bipy)_2]^{3+}$

oxidation process is most likely to proceed via inner sphere mechanism.

BIBLIOGRAPHY

- El-Barghouthi, M. I.; Al-Sou'od, K. A.; Abu-Nawas, A. H.; Abu-El-Halawah, R. Int J Chem 2004, 14, 77– 84.
- Zmachinskii, E.; Malishevskaya, L. I. J Gen Chem 1937, 7, 2693–2695.
- 3. Rigby, W. J Chem Soc, Perkin Trans II 1951, 793–795.
- Jarrar, A.; El-Zaru, R. A. J Chem Educ 1977, 54, 326– 327.
- Weiss, M.; Appel, M. J Am Chem Soc 1932, 70, 3666– 3670.
- Yamarnoto, J.; Murakami, M.; Kameoka, N. Bull Chem Soc Jpn 1982, 55, 345–346.
- Hatanaka, Y.; Imamoto, T.; Yokoyama, M. Tetrahedron Lett 1983, 24, 2399–2400.
- Firouzabadi, H.; Iranpoor, N.; Kiaeezadeh, F.; Toofan, J. Tetrahedron 1986, 42, 719–725.
- Binoy, J.; Vishnu, M. V.; Sreedharan, P.; Vadakkan, J. Synth Commun 2002, 32, 2495–2497.
- Dey, I.; Nongkynrih, I.; Mahanti, M. K. Oxidation Communications 1993, 16, 124–127.
- Chaplin, R. P.; Walpole, A. S.; Zadro, S.; Vorlow, S.; Wainwright, M. S. J Mol Cat 1984, 22, 269–271.
- 12. Belkov, I. A.; Kuzbas, F. G. Kinet I Katal 1988, 29, 1038–1040.
- 13. Hodali, H. A.; El-Zaru, R. A. Tetrahedron 1990, 9, 2299–2304.

- Sun, W.; Norikazu, U.; Nakamura, A. Tetrahedron 1993, 49, 1357–1370.
- Rao, P. V. S.; Suryanarayana, M.; Subbaiah, K. V.; Murty, P. S. N. J Inorg Nucl Chem 1980, 42, 1181–1184.
- Subbaiah, K. V.; Rao, P. V. S.; Murty, B. A. N.; Murty, P. S. N. J. Indian Chem Soc 1979, 56, 1213–1216.
- 17. Pelizzetti, E.; Mentasi, E.; Pramauro, E. Inorg Chem 1976, 15, 2898–2900.
- Sharaydeh, B. F.; Abu-Eid, M. A.; Zatar, N. A.; Abu-Obied, A.; Khamis, M.; Kanan, K. Sci Technol 1994, 22, 355–358.
- Frank, M. S.; Ramaiah, A. K.; Rao, P. V. K. Ind J Chem, Sect A 1979, 18, 369–371.
- 20. Pelizzetti, E.; Mentasi, E.; Pramauro, E. Inorg Chem 1978, 17, 1688–1690.
- 21. Lutz, R. E.; Murphey, R. S. J. J Am Chem Soc 1949, 71, 478–481.
- Mubarak, M. S. Kinetics of Oxidation of Benzoin by Hexacyanoferrate(III) in Alkaline Medium; Mubarak, M. S, (Ed.); University of Jordan: Amman, Jordan, 1977.
- Yousef, A. A. Solvent Effects on the Oxidation of Benzoins by Hexacyanoferrate(III); Yousef, A. A, (Ed.); University of Jordan: Amman, Jordan, 1989.
- 24. Nóberga, J. A.; Rocha, F. R. P. J Chem Educ 1997, 74, 560–562.
- 25. Baxendale, J. H.; George, P. Trans Faraday Soc 1950, 46, 55–63.
- Thompson, J. C.; Mottola, H. A. Anal Chem 1984, 56, 755–757.
- Milburn, R. M.; Vosburgh, W. C. J Am Chem Soc 1955, 77, 1352–1358.