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Kinetics and Mechanism of Ruthenium(III) Catalyzed Oxidation of Chloromphenicol—An Antibiotic Drug by Diperiodatocuprate(III) in Aqueous Alkaline Medium[1]

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Abstract—The kinetics of ruthenium(III) catalyzed oxidation of chloramphenicol (CHP) by diperiodatocuprate(III) (DPC) in aqueous alkaline medium at a constant ionic strength of 0.1 mol l^{-1} was studied spectrophotometrically. The reaction between DPC and CHP in alkaline medium exhibits 1 : 2 stoichiometry (CHP : DPC). The main oxidation products were identified by spot test, IR, NMR, and GC-MS spectral studies. The reaction is first order with respect to ruthenium(III) and DPC concentrations. The order with respect to chloramphenicol concentration varies from first order to zero order as the chloramphenicol concentration increases. As the alkali concentration increases the reaction rate increases with fractional order dependence on alkali concentration. Increase in periodate concentration decreases the rate. A mechanism adequately describing the observed regularities is proposed. The reaction constants involved in the different steps of the mechanism were calculated. The activation parameters with respect to limiting step of the mechanism are computed and discussed. Thermodynamic quantities are determined.

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In recent years the study of highest oxidation states of transition metals has intrigued many researchers. Transition metals in a higher oxidation state can be stabilized by chelating with suitable polydentate ligands. Metal chelates such as diperiodatocuprate(III) [1], diperiodatoargentate(III) [2], diperiodatonickelate(IV) [3] are good oxidants in a medium with an appropriate pH value. Periodate and tellurate complexes of copper in its trivalent state have been extensively used in the analysis of several organic compounds [4]. The kinetics of self-decomposition of these complexes was studied in detail [5]. Copper(III) is shown to be an intermediate in the copper(II)-catalyzed oxidation of amino acids by peroxydisulphate [6]. The oxidation reaction usually involves the copper(II)–copper(I) couple, and such aspects are dealt in different reviews [7, 8]. The use of diperiodatocuprate(III) (**DPC**) as an oxidant in alkaline medium is new and restricted to a few cases due to its limited solubility in aqueous medium. DPC is a versatile one-electron oxidant for various organic compounds in alkaline medium and its use as an analytical reagent is well recognized [9]. Copper complexes have occupied a major place in oxidation chemistry due to their abundance and relevance in biological chemistry [10–14]. Copper(III) is involved in many biological electron transfer reactions [15]. When copper(III) periodate complex is an oxidant, multiple equilibrium steps between different copper(III) species are involved, and it would be of interest to know, which of the species is the active oxidant.

Ruthenium(III) is an efficient catalyst in many redox reactions [16] due to its ability to form intermediate complexes, free radicals, and multiple oxidation state of ruthenium [16, 17]. Chloramphenicol (**CHP**) is effective against a wide variety of Gram-positive and Gram-negative bacteria including most anaerobic organisms. It is considered as a prototypical broad-band spectrum antibiotic, alongside the tetracyclines. The clinical application of chloramphenicol was in the treatment of typhoid. In view of potential pharmaceutical importance of chloramphenicol and lack of literature on the oxidation of this drug by any oxidant and the complexity of the reaction, a detailed study of its oxidation becomes important. Here, the results on the CHP oxidation by DPC is presented, and monoperiodatocuprate (**MPC**) is understood as an active intermediate.

EXPERIMENTAL

All chemicals used were of A.R. grade, and double distilled water was used. The solution of chloramphenicol ("Sisco Chem.") was prepared by dissolving known amount of the samples in distilled water. The purity of the samples was checked by their melting point (150°C). (Literarure: 149–153°C). Ruthenium(III) stock solution is made by dissolving a known weight portion of $RuCl_3$ ("SD Fine-Chem

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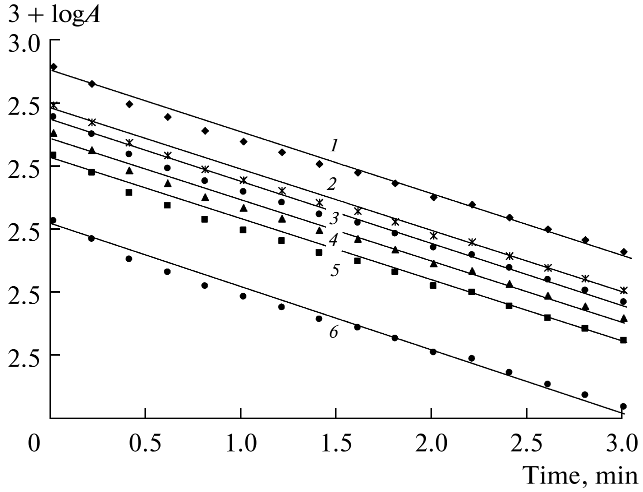


Fig. 1. First order plots of oxidation of chloramphenicol by DPC at 25°C, [CHP] = 5.0×10^{-4} mol/l, [OH^-]] = 0.05 mol/l, [Ru(III)] = 5.0×10^{-7} mol/l, and I = 0.10 mol/l; [DPC] $\times 10^5$ mol/l: 0.5 (*1*), 1.0 (*2*), 2.0 (*3*), 3.0 (*4*), 4.0 (*5*) and 5.0 (*6*).

Ltd.", India) in 0.20 mol/l HCl. Mercury is added to ruthenium(III) solution to reduce any ruthenium(IV) formed during the preparation of ruthenium(III) stock solution and kept for a day. The ruthenium(III) concentration is assayed by EDTA titration [18]. Dilute solution of ruthenium(III) are made from the stock solution as required. The copper(III) periodato complex was prepared [19] and standardized by a standard procedure [20]. The copper(II) solution was made by dissolving the known amount of copper sulphate ("BDH", India) in distilled water. Periodate solution was prepared by dissolving the required amount of sample in hot water and used after 24 h. Its concentration was ascertained iodometrically [21] at neutral pH adjusted by a phosphate buffer. KOH and KNO_3 solutions were employed to maintain the required alkalinity and ionic strength, respectively.

Instruments Used

For kinetic measurements, a Peltier Accessory (temperature control) attached to Varian CARY 50 Bio UV-visible spectrophotometer ("Varian", Victoria-3170, Australia) connected to a rapid kinetic accessory SFA-12 ("Hi-Tech Scientific Ltd.", U.K.) was used.

For product analysis, a chromato-mass spectrometer QP-2010S ("Shimadzu"), Nicolet 5700 FT-IR spectrometer ("Thermo", USA), 300 MHz ^{1}H NMR spectrophotometer ("Bruker", Switzerland) were used.

For pH measurements, an LI-120 pH-meter ("Elico Ltd.", India) was used.

The ruthenium(III) catalyzed oxidation of CHP by DPC was followed under pseudo-first order conditions at [CHP] > [DPC] at 25 ± 0.1°C. The reaction was initiated by adding DPC to chloramphenicol which also contained the required concentration of ruthenium(III), KNO_3, KOH and KIO_4. Since the initial reaction was too fast to be monitored by usual methods, the course of the reaction was followed by monitoring the decrease in the DPC absorbance at its absorption maximum of 415 nm as a function of time in a quartz cell with an optical path length of 1 cm placed in a thermally controlled compartment of a Varian Cary-50 Bio UV-vis spectrophotometer connected to a rapid kinetic accessory (SFA-12, "Hi-Tech"). Earlier the extinction coefficient was determined at 415 nm for different concentrations ($\varepsilon = 6235 \pm 250$ l mol^{-1} cm^{-1}). It was verified that there is negligible interference from other species present in the reaction mixture at this wavelength.

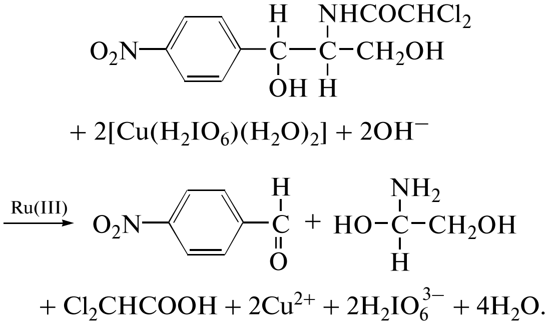
The pseudo-first order rate constants, k_{obs}, were determined from log*A* vs. *t* plots (Fig. 1), where *A* is the absorbance at 415 nm, and *t* is the reaction time. The plots were linear up to 80% completion of the reaction. The rate constants were reproducible within ±5% and are the average of at least three independent kinetic runs. In the kinetic experiments, a constant concentration of KIO_4 (5.0×10^{-5} mol/l) was used throughout the study unless otherwise stated, and excess periodate was present in DPC. The possibility of the chloramphenicol oxidation by periodate in alkaline medium at 25°C was verified. It was found that there is no significant reaction between CHP and periodate under the experimental conditions employed compared to the DPC oxidation of chloramphenicol. The concentration of periodate and OH^- was calculated by considering the amount present in the DPC solution and that additionally added. Kinetic runs were also carried out in N_2 atmosphere in order to understand the effect of dissolved oxygen on the reaction. No significant difference in the results was obtained in the presence and in the absence of nitrogen.

In view of ubiquitous contamination with carbonate in the basic medium, the effect of carbonate was also studied. Added carbonate has no effect on the reaction rates. The spectral changes at 415 nm during the reaction show (Fig. 2) that the concentration of DPC decreases.

RESULTS AND DISCUSSION

Stoichiometry and Product Analysis

Different sets of reaction mixtures containing excess of DPC to chloramphenicol in the presence of constant amount of OH^-, ruthenium(III), KNO_3, KIO_4 were kept for 6 h in a closed vessel under inert atmosphere. The remaining DPC concentration was estimated by spectrophotometrically at 415 nm. The results, indicated that one mole of CHP requires two moles of DPC as shown in Eq. (I).



(I)

After completion of the reaction, the reaction mixture was acidified, concentrated, and extracted with ether. The ether layer was subjected to column chromatography [22], and the fractions were subjected to spectral investigations. From the IR, GC-MS, and ^1H NMR spectra, the main oxidation product was identified as *p*-nitrobenzaldehyde. IR spectrum showed C=O stretching mode for aldehyde functional group at 1708 cm^{-1}, while $-NO_2$ stretching mode was observed at 1349 cm^{-1}. The presence of *p*-nitrobenzaldehyde was also confirmed by GC-MS analysis obtained on a Shimadzu's 17A gas chromatograph with a Shimadzu's XP-5000A mass spectrometer using electron impact (EI) ionization technique. The mass spectrum showed the base peak at 151 amu consistent with the molecular ion of 151 amu. All other peaks observed in the GC-MS can be interpreted in accordance with the structure of *p*-nitrobenzaldehyde. Further *p*-nitrobenzaldehyde was characterized by its ^1H NMR spectrum (DMSO): 10.17 (s, 1H, CHO), 8.18 (d, 2H, ArH), 7.67 (d, 2H, ArH). Another product 2-amino-2-hydroxyethanol was confirmed by GC-MS spectrum, which showed the molecular ion peak at 77 amu. Dichloroacetic acid was confirmed by spot test [23].

Reaction Orders

The reaction orders were determined from the slopes of the k_{obs} *vs.* concentration plots in logarithmic coordinates by varying the concentrations of chloramphenicol, alkali, ruthenium(III) and periodate in turn while keeping all other concentrations and conditions constant.

The oxidant DPC concentrations were varied in the range of 5.0×10^{-6} to 5.0×10^{-5} mol/l, and fairly constant k_{obs} values indicated that order with respect to DPC concentration was unity (Table 1). This was also confirmed by linearity and almost parallel plots of log*A* vs. *t* up to 80% completion of the reaction as shown in Fig. 1.

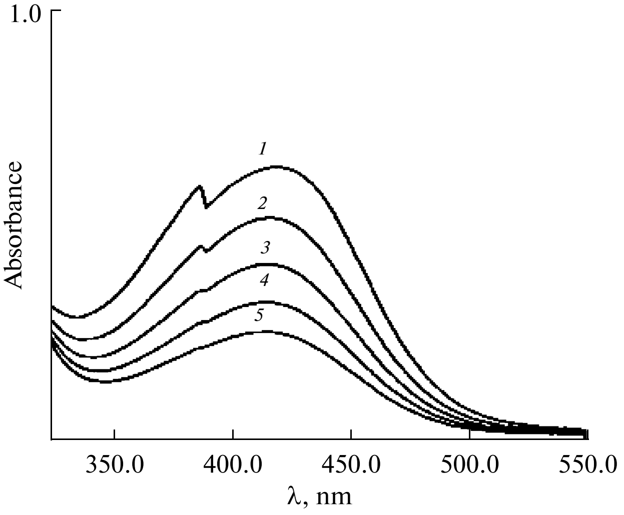


Fig. 2. Spectral changes during the oxidation of chloramphenicol by DPC at 25°C, [DPC] = 5.0×10^5 mol/l, [CHP] = 5.0×10^{-4} mol/l, [OH^-] = 0.05 mol/l, [Ru(III)] = 5.0×10^{-7} mol/l, and $I = 0.10$ mol/l (scanning time interval is 1.0 min).

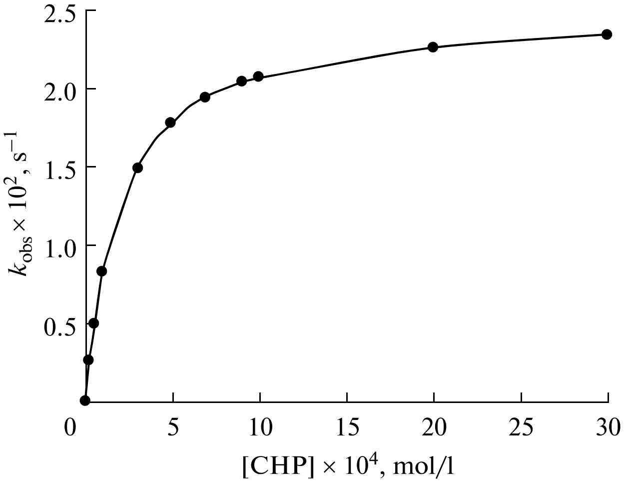


Fig. 3. Dependence of k_{obs} on [CHP].

The effect of chloramphenicol on the reaction rate was studied at constant concentrations of alkali, DPC, and periodate at the constant ionic strength of 0.10 mol/l. The CHP concentration was varied in the range of 0.25×10^{-4} to 3.0×10^{-3} mol/l. The k_{obs} values increased with an increase in the concentration of CHP (Table 1). It is observed that, as the CHP concentration increases, the rate of reaction increases with saturation: at [CHP] $\geq 2 \times 10^{-3}$ mol/l, k_{obs} almost does not change, and the order with respect to CHP concentration has changed from the first to zero order (Fig. 3). In the lower concentration range, from 0.25×10^{-4} to 1.0×10^{-4} mol/l, the rate of oxidation shows first-order dependence on [CHP]. At intermediate concentrations of CHP, 1.0×10^{-4} to 5.0×10^{-4} mol/l, the reaction order decreases, and, at the high concentrations $\geq 2.0 \times 10^{-3}$ mol/l, the rate of oxidation is independent of [CHP].

The effect of alkali concentration on the reaction was studied in the range 0.01 to 0.1 mol/l at constant

Table 1. Effect of variation of [DPC], [CHP], [Ru(III)], [IO_4^-] and [OH^-] on the oxidation of chloramphenicol by DPC at 25°C and $I = 0.10$ mol/l

[DPC] × 10^5, mol/l	[CHP] × 10^4, mol/l	[OH^-], mol/l	[IO_4^-] × 10^5, mol/l	[Ru(III)] ×10^7, mol/l	k_{obs} ×10^2, s^{-1}
0.5	5.0	0.05	5.0	5.0	1.75
1.0	5.0	0.05	5.0	5.0	1.71
2.0	5.0	0.05	5.0	5.0	1.75
3.0	5.0	0.05	5.0	5.0	1.76
4.0	5.0	0.05	5.0	5.0	1.79
5.0	5.0	0.05	5.0	5.0	1.78
5.0	0.25	0.05	5.0	5.0	0.26
5.0	0.50	0.05	5.0	5.0	0.50
5.0	1.00	0.05	5.0	5.0	0.83
5.0	3.00	0.05	5.0	5.0	1.49
5.0	5.00	0.05	5.0	5.0	1.78
5.0	7.00	0.05	5.0	5.0	1.94
5.0	9.00	0.05	5.0	5.0	2.04
5.0	10.0	0.05	5.0	5.0	2.07
5.0	20.0	0.05	5.0	5.0	2.26
5.0	30.0	0.05	5.0	5.0	2.34
5.0	5.0	0.01	5.0	5.0	1.41
5.0	5.0	0.03	5.0	5.0	1.70
5.0	5.0	0.05	5.0	5.0	1.78
5.0	5.0	0.07	5.0	5.0	1.81
5.0	5.0	0.09	5.0	5.0	1.83
5.0	5.0	0.10	5.0	5.0	1.84
5.0	5.0	0.05	1.0	5.0	2.75
5.0	5.0	0.05	3.0	5.0	2.23
5.0	5.0	0.05	5.0	5.0	1.78
5.0	5.0	0.05	7.0	5.0	1.48
5.0	5.0	0.05	9.0	5.0	1.30
5.0	5.0	0.05	10.0	5.0	1.21
5.0	5.0	0.05	5.0	0.5	0.18
5.0	5.0	0.05	5.0	1.0	0.36
5.0	5.0	0.05	5.0	2.0	0.71
5.0	5.0	0.05	5.0	3.0	1.07
5.0	5.0	0.05	5.0	4.0	1.42
5.0	5.0	0.05	5.0	5.0	1.78

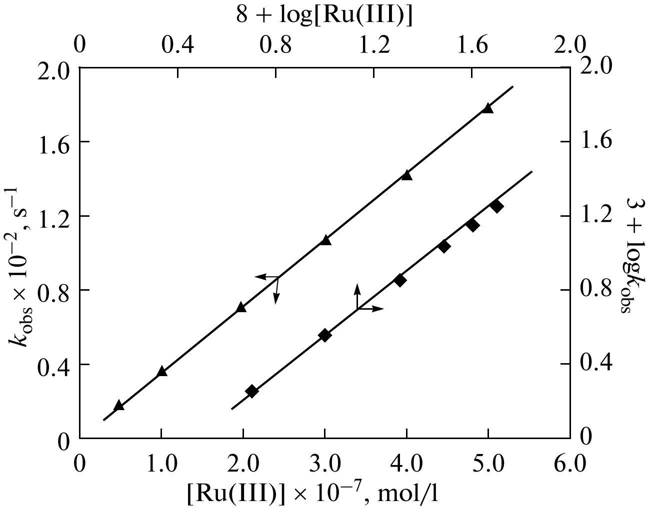


Fig. 4. Effect of [Ru(III)] catalyst on the oxidation of CHP by DPC in aqueous alkaline medium at 25°C, [DPC] = 5.0×10^5 mol/l, [CHP] = 5.0×10^{-4} mol/l, [OH^-] = 0.05 mol/l, and $I = 0.10$ mol/l.

concentrations of CPH, DPC, and periodate at constant ionic strength of 0.10 mol/l at 25°C. The rate constants increased with an increase in the alkali concentration (Table 1). The order was found to be less than unity (0.31).

The effect of periodate was studied by varying the periodate concentration from 1.0×10^{-5} to 1.0×10^{-4} mol/l keeping all other reactant concentrations constant. It was found that the added periodate had a retarding effect on the reaction rate (Table 1).

At constant oxidant, reductant, and alkali concentration of 5.0×10^{-5}, 5.0×10^{-4} and 0.05 mol/l, respectively, and $I = 0.10$ mol/l, the catalyst ruthenium(III) concentration was varied between 1.0×10^{-7} and 1.0×10^{-6} mol/l (Table 1). The order with respect to ruthenium concentration was found to be unity. As catalyst concentration increases the rate of the reaction also increases, and activity of catalyst reaches a maximum after [Ru(III)] $\geq 5.0 \times 10^{-7}$ mol/l under the above mentioned experimental conditions (Fig. 4). Under condition used, the k_{obs} for uncatalyzed reaction is 1.44×10^{-3} s^{-1} and is negligible compared to the mediated reaction (Table 1).

The effect of ionic strength was studied by varying the potassium nitrate concentration from 0.01 to 1.0 mol/l at constant concentration of DPC, ruthenium(III), CPH, periodate, and alkali. It was found that increasing ionic strength had negligible effect on the rate constant.

The effect of dielectric constant was studied by varying the *t*-butanol–water (*v*/*v*) content in the reaction mixture with all the other conditions being kept constant. Decreasing the dielectric constant of the medium had no effect on the rate of the reaction.

The externally added products, *p*-nitrobenzaldehyde and copper(II) ($CuSO_4$) in the concentration range 1.0×10^{-6} to 1.0×10^{-5} mol/l did not have any significant effect on reaction rate.

The participation of free radicals in the reaction was examined as follows. The reaction mixture, to which a known quantity of acrylonitrile monomer was initially added, was kept for 2 h in an inert atmosphere. On diluting the reaction mixture with methanol, a white precipitate was formed indicating the participation of free radicals. The blank experiments of either DPC or Ru(III) alone with acrylonitrile did not induce any polymerization under the same conditions as those induced for reaction mixture. Initially added acrylonitrile decreased the reaction rate also indicating free radical intervention.

The kinetics was studied at four different temperatures under varying concentrations of CHP, alkali, and periodate, keeping other conditions constant. The

Table 2. Thermodynamic activation parameters for the Ru(III)-catalyzed oxidation of CHP by DPC in aqueous alkaline medium with respect to the slow step of the scheme

E_a for k, kJ mol^{-1}	$\Delta H^\#$ for k and ΔH for K_i, kJ/mol	$\Delta S^\#$ for k and ΔS for K_i, J K^{-1} mol^{-1}	$\Delta G^\#$ for k and ΔG for K_i, kJ/mol
14 ± 0.5	12 ± 0.2	−114 ± 4	46 ± 2.4
–	26 ± 2	161 ± 8	−23 ± 0.5
–	−29 ± 0.8	−135 ± 6	11.7 ± 0.2
–	26 ± 0.3	161 ± 5	−23 ± 0.4

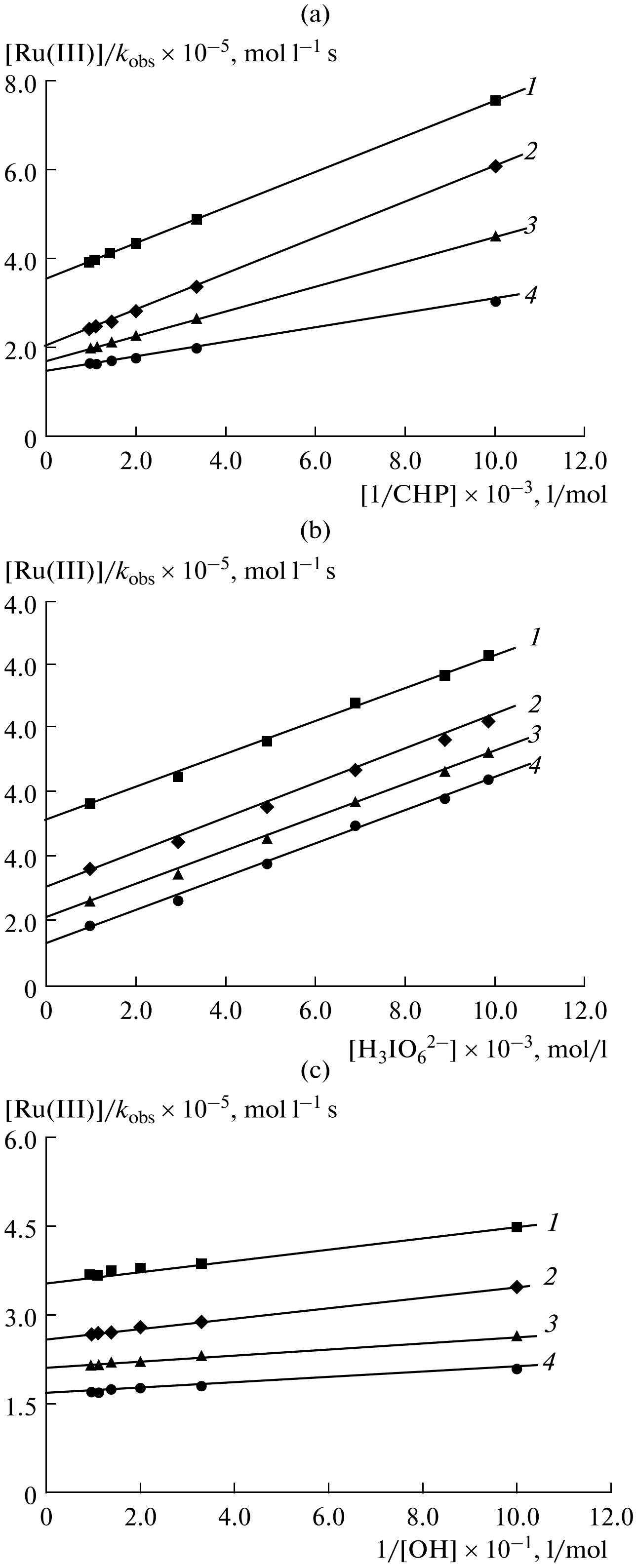


Fig. 5. Plots of [Ru(III)]/k_{obs} *vs.* 1/[CHP] (a), [$H_3IO_6^{2-}$] (b), and 1/[OH^-] (c) at different temperatures, K: 288 (*1*), 298 (*2*), 308 (*3*), and 318 (*4*) (conditions as in Table 1).

reaction rate was found to increase with an increase in temperature. The rate constants (k) of the slow step of Scheme were obtained from the slopes and intercepts of [Ru(III)]/k_{obs} vs. 1/[OH^-], [Ru(III)]/k_{obs} vs. 1/[CHP], and Ru(III)]/k_{obs} vs. [$H_3IO_6^{2-}$] plots (Figs. 5a–5c) at four different temperatures. The energy of activation corresponding to these rate constants was evaluated from the Arrhenius plot of log$k$ vs. $1/T$ from which otner activation parameters were calculated (Table 2).

It has been pointed out by Moelwyn–Hughes [24] that, in presence of the catalyst, the uncatalyzed and catalyzed reactions proceed simultaneously, so that

$$k_{obs} = k_u + K_C[\text{catalyst}]^x. \quad (1)$$

Here k_{obs} is the observed pseudo-first order rate constant in the presence of ruthenium(III) catalyst, k_u is the observed pseudo-first order rate constant for uncatalyzed reaction, K_C is the catalytic constant and "x" is the order of the reaction with respect to [Ru(III)]. In the present investigation, x value is found to be unity. Then the value of K_C can be calculated from Eq. (2)

$$K_C = (k_{obs} - k_u)/[\text{catalyst}] = k_C/[\text{catalyst}], \quad (2)$$

where $k_C = k_{obs} - k_u$. For the present experimental condition and for the standard kinetic run, the K_C was found to be 3.26×10^4 l mol^{-1} s^{-1} at 298 K.

The water soluble copper(III) periodate complex $[Cu(HIO_6)_2(OH)_2]^{7-}$ is reported [25]. However, in an aqueous alkaline medium and at high pH range employed in the study, periodate is unlikely to exist as HIO_6^{4-} that is evident from its involvement in the multiple equilibria (II)–(IV) [26–28], depending on the pH of the solution.

$$H_5IO_6 \underset{}{\overset{K_1}{\rightleftharpoons}} H_4IO_6^- + H^+ \quad K_1 = 5.1 \times 10^{-4}, \quad (II)$$

$$H_4IO_6^- \overset{K_2}{\rightleftharpoons} H_3IO_6^{2-} + H^+ \quad K_2 = 4.9 \times 10^{-9}, \quad (III)$$

$$H_4IO_6^{2-} \overset{K_3}{\rightleftharpoons} H_2IO_6^{3-} + H^+ \quad K_3 = 2.5 \times 10^{-12}. \quad (IV)$$

Periodic acid exists as H_5IO_6 in acid medium and as $H_4IO_6^-$ near pH 7. Thus, under the conditions employed in alkaline medium, main species are expected to be $H_3IO_6^{2-}$ and $H_2IO_6^{3-}$. Thus, at pH employed in this study, the soluble copper(III) periodate complex exist as diperiodatocuprate(III), $[Cu(OH)_2(H_3IO_6)_2]^{2-}$, a conclusion also supported by the literature [27, 28].

The reaction between the diperiodatocuprate(III) complex and CHP in alkaline medium has stoichiometry 1 : 2 (CHP : DPC) with first order dependence on the DPC and ruthemium(III) concentrations and less than the first order in CHP and alkali concentrations, a negative fractional order dependence on the periodate concentration. No effect of added products was observed. Based on the experimental results, a mechanism is proposed for which all the observed orders in respect to each constituent such as [oxidant], [reductant], [Ru^{III}], [OH^-] and [IO_4^-] may be well accommodated. Lister [29] proposed three forms of copper(III) periodate in alkaline medium as diperiodatocuprate(III) (DPC), monoperiodatocuprate(III) (MPC) and tetraperiodatocuprate(III). The latter is ruled out, as its equilibrium constant is 8.0×10^{-11} mol/l at 40°C. In the present kinetic study, DPC and MPC are to be considered as active forms of copper(III) periodate complex. It may be expected that lower periodate complex such as MPC is more important in the reaction than DPC. The increase in the rate with an increase in alkali concentration suggests that equilibia of copper(III) periodate complexes are possible as in Eqs. (V) and (VI).

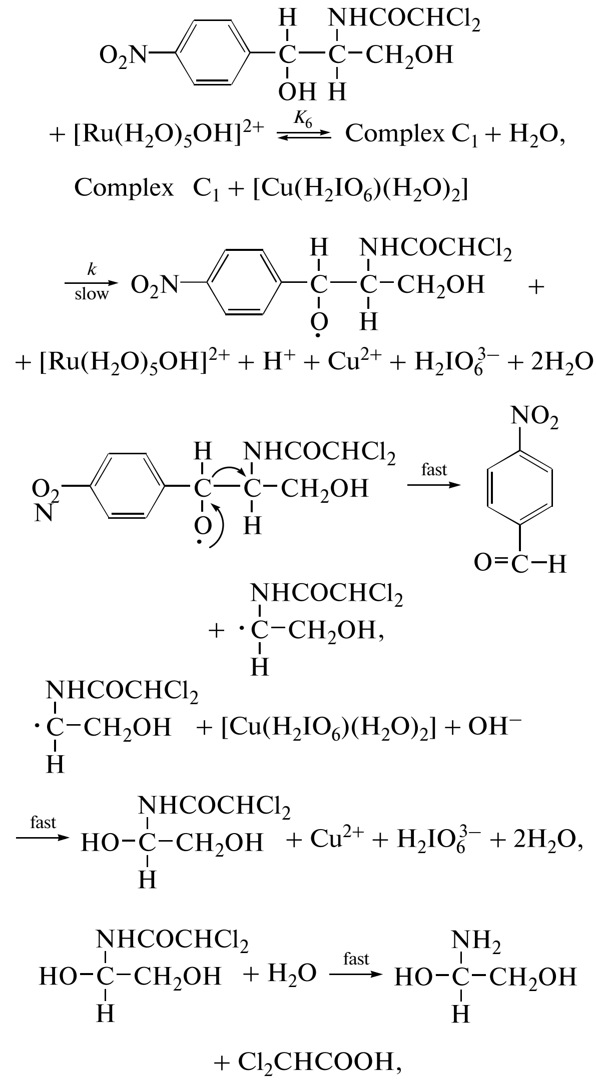
$$[Cu(H_3IO_6)_2]^- + [OH^-] \overset{K_4}{\rightleftharpoons} [Cu(H_2IO_6)(H_3IO_6)]^{2-} + H_2O, \quad (V)$$

$$[Cu(H_2IO_6)(H_3IO_6)]^{2-} + 2H_2O \overset{K_5}{\rightleftharpoons} [Cu(H_2IO_6)(H_2O)_2] + [H_3IO_6]^{2-}. \quad (VI)$$

The reaction order less than unit in respect to [CHP] presumably results from the formation of a complex C_1 between the ruthenium(III) species and CHP prior to the formation of the products. The ruthenium(III) catalyst forms the complex with the substrate, which enhances the reducing property of the substrate. This complex C_1 reacts with one molecule of MPC in the fast step to give a free radical intermediate species from CHP, Cu(III) being reduced to Cu(II) by accepting an electron from O–H bond cleavage and ruthenium(III) being regenerated. The C–C bond in the free radical cleaves in a subsequent fast step to give *p*-nitrobenzaldehyde and another free radical. In the next fast step, the free radical thus formed reacts with another molecule of MPC in the presence of OH^- to form 2,2-dichloro-N-(1,2-dihydroxyethyl)acetamide. In the further fast steps 2,2-dichloro-N-(1,2-dihydroxyethyl)acetamide undergoes hydrolysis to give another product 2-amino~2-hydroxyethanol and dichloroacetic acid. The detailed mechanism for the oxidation of CHP by diperiodatocuprate(III) is presented in scheme.

$$[Cu(H_3IO_6)_2]^- + [OH^-] \overset{K_4}{\rightleftharpoons} [Cu(H_2IO_6)(H_3IO_6)]^{2-} + H_2O,$$

$$[Cu(H_2IO_6)(H_3IO_6)]^{2-} + 2H_2O \overset{K_5}{\rightleftharpoons} [Cu(H_2IO_6)(H_2O)_2] + [H_3IO_6]^{2-},$$



$$+ [Ru(H_2O)_5OH]^{2+} \overset{K_6}{\rightleftharpoons} \text{Complex } C_1 + H_2O,$$

$$\text{Complex } C_1 + [Cu(H_2IO_6)(H_2O)_2]$$

$$\xrightarrow[\text{slow}]{k} \ldots + [Ru(H_2O)_5OH]^{2+} + H^+ + Cu^{2+} + H_2IO_6^{3-} + 2H_2O,$$

$$+ \cdot\overset{NHCOCHCl_2}{\underset{H}{C}}-CH_2OH,$$

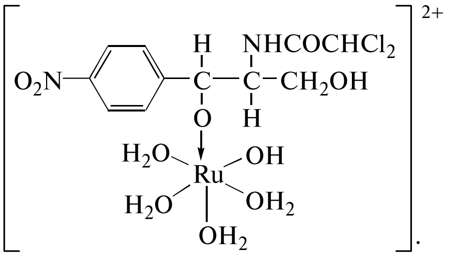
$$\cdot\overset{NHCOCHCl_2}{\underset{H}{C}}-CH_2OH + [Cu(H_2IO_6)(H_2O)_2] + OH^- \xrightarrow{\text{fast}} HO-\overset{NHCOCHCl_2}{\underset{H}{C}}-CH_2OH + Cu^{2+} + H_2IO_6^{3-} + 2H_2O,$$

$$HO-\overset{NHCOCHCl_2}{\underset{H}{C}}-CH_2OH + H_2O \xrightarrow{\text{fast}} HO-\overset{NH_2}{\underset{H}{C}}-CH_2OH + Cl_2CHCOOH,$$

$$H^+ + OH^- \xrightarrow{\text{fast}} H_2O.$$

Scheme.

The probable structure of the complex C_1 is given by:



Complex C_1 shows that hydroxyl group oxygen was coordinated to the metal complex. Since Scheme is in accordance with the generally well accepted principle of non-complementary oxidations taking place in sequence of one-electron steps, the reaction between substrate and oxidant would afford a radical intermediate. A free radical scavenging experiment revealed the possibility. This type of radical intermediate has also been observed in earlier work [30]. A bathochromic shift of about 6 nm from 271 to 277 nm in the spectra of CHP observed confirms the complex formation between ruthenium(III) and CHP in their mixture. The Michaelis–Menten plot also proves the complex formation between the catalyst and substrate and explains the complex dependence on [CHP]. Such type of complex formation between a substrate and a catalyst has been observed in literature [31, 32]. The rate law (11) for the intermediate concentrations of DPC can be derived from scheme as follows. According to scheme,

$$\text{Rate} = -\frac{d[DPC]}{dt} = k[\mathrm{Cu(H_2IO_6)(H_2O)_2[C_1]}$$

$$= \frac{kK_5K_6[\mathrm{Cu(H_2IO_6)(H_3IO_6)}]^{2-}[\mathrm{CHP}][\mathrm{Ru(III)}]}{[\mathrm{H_3IO_6^{2-}}]} \qquad (3)$$

$$= \frac{kK_4K_5K_6[\mathrm{CHP}][\mathrm{Ru(III)}][\mathrm{DPC}]_f[\mathrm{OH^-}]_f}{[\mathrm{H_3IO_6^{2-}}]},$$

where $[\mathrm{DPC}]_t$ and $[\mathrm{DPC}]_f$ refer to total and free DPC concentrations. The total concentration of DPC, $[\mathrm{DPC}]_t$,is given by

$$[\mathrm{DPC}]_t = [\mathrm{DPC}]_f + [\mathrm{Cu(H_2IO_6)(H_2IO_6)}]^{2-} + [\mathrm{Cu(H_2IO_6)(H_2O)_2}] = [\mathrm{DPC}]_f \times \left\{\frac{[\mathrm{H_3IO_6^{2-}}] + K_4[\mathrm{H_3IO_6^{2-}}][\mathrm{OH^-}] + K_4K_5[\mathrm{OH^-}]}{[\mathrm{H_3IO_6^{2-}}]}\right\}. \qquad (4)$$

Therefore $[\mathrm{DPC}]_f$ is given by

$$[\mathrm{DPC}]_f = \frac{[\mathrm{DPC}]_t[\mathrm{H_3IO_6^{2-}}]}{[\mathrm{H_3IO_6^{2-}}] + K_4[\mathrm{H_3IO_6^{2-}}][\mathrm{OH^-}] + K_4K_5[\mathrm{OH^-}]}. \qquad (5)$$

Since concentrations of ruthenium(III) used are several orders of magnitude lower than [CHP],

$$[\mathrm{CHP}]_t = [\mathrm{CHP}]_f. \qquad (6)$$

The free [Ru(IIl)] can be written as

$$[\mathrm{Ru(III)}]_f = [\mathrm{Ru(III)}]_f = \frac{[\mathrm{Ru(III)}]_f}{1 + K_6[\mathrm{CHP}]}. \qquad (7)$$

In the same manner, due to the low concentrations of [DPC] and $[\mathrm{H_3IO_6^{2-}}]$ in comparison to $[\mathrm{OH^-}]$

$$[\mathrm{OH^-}]_f = [\mathrm{OH^-}]_t. \qquad (8)$$

Substituting Eqs. (5)–(8) in Eq. (3) and omitting t and f, we get

$$\text{Rate} = \frac{kK_4K_5K_6[\mathrm{CHP}][\mathrm{DPC}][\mathrm{OH^-}][\mathrm{Ru(III)}]}{[\mathrm{H_3IO_6^{2-}}] + K_4[\mathrm{H_3IO_6^{2-}}][\mathrm{OH^-}] + K_4K_5[\mathrm{OH^-}] + K_4K_5K_6[\mathrm{OH^-}][\mathrm{CHP}]}, \quad \text{or} \qquad (9)$$

$$\frac{\text{Rate}}{[\mathrm{DPC}]} = k_{\mathrm{obs}} = \frac{kK_4K_5K_6[\mathrm{CHP}][\mathrm{DPC}][\mathrm{OH^-}][\mathrm{Ru(III)}]}{[\mathrm{H_3IO_6^{2-}}] + K_4[\mathrm{H_3IO_6^{2-}}][\mathrm{OH^-}] + K_4K_5[\mathrm{OH^-}] + K_4K_5K_6[\mathrm{OH^-}][\mathrm{CHP}]}. \qquad (10)$$

The rate law (10) can be transformed to the following form, which is suitable for verification.

$$\frac{[\mathrm{Ru(III)}]}{k_{\mathrm{obs}}} = \frac{[\mathrm{H_3IO_6^{2-}}]}{kK_4K_5K_6[\mathrm{OH^-}][\mathrm{CHP}]} + \frac{[\mathrm{H_3IO_6^{2-}}]}{kK_5K_6[\mathrm{CHP}]} + \frac{1}{kK_6[\mathrm{CHP}]} + \frac{1}{k}. \qquad (11)$$

According to Eq. (11), other conditions being constant, plots of$[\mathrm{Ru(III)}]/k_{\mathrm{obs}}$ vs. $1/[\mathrm{OH^-}]$, $[\mathrm{Ru(III)}]/k_{\mathrm{obs}}$ vs. 1/[CHP], and $\mathrm{Ru(III)}]/k_{\mathrm{obs}}$ vs. $[\mathrm{H_3IO_6^{2-}}]$ should be linear and found to be so (Fig. 5). From the slopes and intercepts of these plots the following values of K_4, K_5, K_6 and k at 25°C were calculated: 2.11×10^{-1} 1 mol^{-1}, 1.62×10^{-2} mol/l, 6.43×10^{3} 1 mol^{-1} s^{-1}, and 4.98×10^{4} s^{-1}, respectively. The values of K_4, K_5 and K_6 are in good agreement with those given in the literature [33]. The equilibrium constant K_4 is far greater than K_5. This may be attributed to the greater tendency of DPC to undergo hydrolysis compared to the dissociation of hydrolyzed species in alkaline medium.

The thermodynamic quantities for the different equilibrium steps of the scheme can be evaluated as follows. The [CHP] and $[\mathrm{OH^-}]$ (as in Table 1) were varied at four different temperatures. The plots of $[\mathrm{Ru(III)}]/k_{\mathrm{obs}}$ vs. $1/[\mathrm{OH^-}]$, $[\mathrm{Ru(III)}]/k_{\mathrm{obs}}$ vs. 1/[CHP], and $[\mathrm{Ru(III)}]/k_{\mathrm{obs}}$ vs. $[\mathrm{H_3IO_6^{2-}}]$ are linear. From the slopes and intercepts, the values of K_4 were calculated at different temperatures, and these values are given in Table 3. vant Hoff dependences ($\log K_4$ vs. $1/T$) were plotted, and the values of the enthalpy of reaction ΔH, the entropy of reaction ΔS, and the free energy of reaction ΔG were calculated (Table 2). A comparison of the these values with those obtained from the slow step of

Table 3. The values of K_4, K_5 and K_6 calculated at different temperatures

Temperature, K	$k \times 10^{-4}$, s^{-1}	K_4, l/mol	$K_5 \times 10^2$, mol/l	$K_6 \times 10^{-3}$, l/mol
288	3.88	0.06 ± 0.01	2.12 ± 0.01	4.1 ± 0.2
298	4.98	0.21 ± 0.01	1.62 ± 0.01	6.3 ± 0.2
308	5.83	0.37 ± 0.01	1.12 ± 0.02	7.7 ± 0.1
318	6.97	0.41 ± 0.02	0.65 ± 0.01	12.3 ± 0.1

the reaction shows that these values mainly refer to the rate limiting step supporting the fact that the reaction before the rate determining step is fairly fast and occurs with the low activation energy [34]. In the same manner, K_5 and K_6 values were calculated at different temperatures (Table 3), and the corresponding values of the thermodynamic parameters are given in Table 2.

The moderate values of $\Delta S^{\#}$ suggest that the intermediate complex is more ordered than the reactants [35]. The value of $\Delta S^{\#}$ within the range of radical reaction has been ascribed to the nature of electron pairing and unpairing processes and to the loss of degrees of freedom formerly available to the reactants upon the formation of rigid transition state [36]. The observed modest enthalpy of activation and relatively low value of the entropy of activation as well as a higher rate constant of slow step indicate that the oxidation presumably occurs via inner-sphere mechanism. The conclusion is supported by earlier observation [37]. The negligible effect of ionic strength explains the involvement of neutral species in scheme. However, it is difficult to interpret the effect of dielectric constant in view of various ions involved. The catalyst ruthenium(III) forms the complex with substrate, which enhances the reducing property of substrate.

Among various species of DPC in alkaline medium, monodiperiodatocuprate(III) (MPC) $(Cu(H_2IO_6)(H_2O)_2)$ is considered as active species for the reaction under study. The active species of ruthenium(III) is found to be $[RuH_2O)_5OH]^{2+}$. The catalytic constants and the activation parameters with reference to catalyst were computed. The results demonstrate that the role of pH in the reaction medium is crucial. Rate constant of the slow step and other equilibrium constants involved in the mechanism evaluated, and activation parameters with respect to slow step of the reaction were computed. The overall mechanistic sequence described here is consistent with product, mechanistic, and kinetic studies.

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