Mechanism of Oxidation of Alanine by Chloroaurate(III) Complexes in Acid Medium: Kinetics of the Rate Processes

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ABSTRACT: The kinetics of the oxidation of alanine by chloroaurate(III) complexes in acetate buffer medium has been investigated. The major oxidation product of alanine has been identified as acetaldehyde by ¹H NMR spectroscopy. Under the experimental conditions, AuCl₄⁻ and AuCl₃(OH)⁻ are the effective oxidizing species of gold(III). The reaction is first order with respect to Au(III) as well as alanine. The effects of H⁺ and Cl⁻ on the second-order rate constant k'_2 have been analyzed, and accordingly the rate law has been deduced: $k'_2 = (k_1[H^+]|Cl^-] + k_3K_4K_5)/(K_4K_5 + [H^+]|Cl^-])$. Increasing dielectric constant of the medium has an accelerating effect on the reaction rate. Activation parameters associated with the overall reaction have been calculated. A mechanism involving the two effective oxidizing species of gold(III) and zwitterionic species of alanine, consistent with the rate law, has been proposed. © 2009 Wiley Periodicals, Inc. Int J Chem Kinet 41: 473–482, 2009

INTRODUCTION

Amino acids are the chemical units or building blocks of protein molecules and enzymes and play a key role in life processes. Amino acids found within protein convey a vast array of chemical versatility. The chemical properties of the amino acids determine the biological activity of the protein. The mechanism of redox reactions involving amino acids is therefore of immense importance. A number of kinetic studies have been reported for the oxidation of some amino acids by different metal ion oxidants such as Fe^{III} [1, 2], Ag^{III} [3, 4], V^V [5, 6], Ce^{IV} [7, 8], Bi^V [9], and Mn^{VII} [10– 14] in aqueous acidic or alkaline medium. Over the past few years studies in the chemistry of gold(I) and gold(II) compounds have received much attention. A few gold(I) compounds are biologically active and are used as anti-inflammatory drugs in the treatment of rheumatoid arthritis [15–18]. There are reports on the kinetic study of the reduction of gold(III) complexes by sulfur^{IV} [19], dimethylsulfide [20], and H₂O₂ [21]. Some recent studies for using gold(III) complexes as

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antitumor and anti-HIV agents have been reported [22-24]. Investigations on redox and ligand exchange reactions of potential gold(I) and gold(III)-cyanide metabolites under biomimetic conditions [25] and those on a metal-centered photo-induced electron transfer reduction of a Au(III)-porphyrin cation in nonpolar solvents [26] have recently been reported. Recently, the reactions of a few gold(III) complexes with serum albumin [27] were investigated using spectroscopic methods and separation techniques. The presence of both weak and strong metal-protein interactions and adduct formation for such gold(III) complexes through coordination at the level of surface histidines has been proposed. In a recent kinetic study in the reduction of gold(III) complexes with calyx(4)areness [28] having thioether groups on their upper rim, both intramolecular and intermolecular mechanisms have been proposed. There is now a growing interest in the production of gold nanoparticles [29], and the oxidation of gold nanoparticles by Au(III)-CTAB complexes has been studied [30]. The mechanism of oxidation depends on whether Au(III) is attached to the micelles. It has been proposed that the CTAB micelles approach the nanoparticles preferentially at the tips, leading to spatially directed oxidation. The kinetics of oxidation of glycolaldehyde [31], a few alkanols [32], aryl alcohols [32], and α -hydroxy acids [33] by Au(III) has already been reported from our laboratory. We have recently studied the kinetics of oxidation of glycine by chloroaurate(III) complexes in acetate buffer [34]. An attempt has therefore been made in the same direction to study whether the oxidation of alanine follows the same mechanistic path. For reasons mentioned earlier [34], the kinetics of the reaction has been studied in the sodium acetate-acetic acid buffer medium.

EXPERIMENTAL

Reagents

All chemicals were of the highest available purity. The solution of DL-alanine (extra pure, AR; SRL, Mumbai, India) was estimated by the formol titration method [35] in the presence of excess neutralized formalin with standard NaOH. Chloroauric acid trihydrate (GR; E. Merck, Darmstadt, Germany) was dissolved in 0.01 mol dm⁻³ HCl, and the strength of the Au(III) stock solution was determined by measurement of the absorbance of the solution (after proper dilution in 1 mol dm⁻³ NaCl and 0.01 mol dm⁻³ HCl) at the absorbance maximum [36] (313 nm, $\varepsilon = 4860 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}$) of AuCl₄⁻. All the solutions were made in doubly distilled water.

Kinetics and Measurements

All the kinetic investigations were studied under pseudo-first-order conditions, where [alanine] and [Cl⁻] were present in large excess compared to [Au^{III}]. Also [H⁺] was kept constant using sodium–acetate– acetic acid buffer. The kinetic progress of the reaction was followed by measuring the absorbance of the reaction mixture at 400 nm for reasons discussed earlier [34]. Alanine and the products are transparent at this wavelength. The kinetic measurements were performed in a Systronics (Ahmedabad, India) visible spectrophotometer (model 106) with a thermostated cell compartment. The plots of log A (A = absorbance) vs. time were linear at least up to two half lives, and the pseudo-first-order rate constants (k_{obs}), evaluated from these plots, were reproducible to within ±5%.

Product Analysis

The reaction mixture (concentration of each reactant being 20 times that under kinetic condition) was allowed to stand for 2 h in a closed vessel and then distilled. The distillate was collected in a closed container, and the following tests were performed with the distillate:

To the first part, five drops of piperidine was added, followed by a few drops of 5% sodium nitroprusside solution and then made alkaline with Na₂CO₃ solution-a blue color developed [37], indicating the presence of CH₃CHO. The second part of the solution was treated with a few drops of sodium nitroprusside solution followed by a bead of NaOH-the development of a red color [38] indicated the presence of CH₃CHO in the reaction product. Another part of the distillate was treated with 2,4-dinitrophenylhydrazine (DNP) solution in 4 N H₂SO₄ when an orange DNP derivative of the product was precipitated. It was filtered, washed thoroughly, and dried, and its melting point was found to be 145°C [39]. ¹H NMR (300 MHz, Bruker DPX 300) of the 2,4-dinitrophenylhydrazone derivative of the reaction product in CDCl₃ displayed a broad signal at δ 11.17 due to N–H; two doublets, one at δ 9.15 (J = 2.3 Hz) and the other at δ 9.1 (J= 2.5 Hz) [total 1H] due to aromatic-H flanked between two $-NO_2$ groups; doublet of doublets at δ 8.34 $(J_1 = 9.9 \text{ Hz}, J_2 = 2.2 \text{ Hz})$ and at $\delta 8.30 (J_1 = 9.9 \text{ Hz},$ $J_2 = 2.4$ Hz) due to aromatic-H ortho to one $-NO_2$ group (total 1H) and doublets at δ 7.97 (J = 9.1 Hz) and δ 7.94 (J = 9.3 Hz) (total 1H) due to aromatic-H meta to both the -NO₂ groups. The NMR picture clearly showed the existence of two isomeric DNP derivatives. Most interestingly, the ¹H NMR spectrum of the said compound furnished a pair of quartets at δ 7.56 (J = 5.3 Hz) and at δ 7.11 (J = 5.5 Hz) along with two doublets at δ 2.14 (J = 5.4 Hz) and the other at δ 2.08 (J = 5.6 Hz). Therefore, it is conclusively proved that two stereoisomeric (svn and anti) 2,4-dinitrophenylhydrazones were present in the DNP derivative of the reaction mixture. For one stereoisomer (presumably *anti*) quartet at δ 7.56 (1H, J = 5.3 Hz) due to olefinic-H and the doublet at δ 2.14 (3H, J = 5.4 Hz) due to the $-CH_3$ group attached with olefinic carbon were furnished. Similarly, for other stereoisomer (presumably syn) quartet at δ 7.11 (1H, J = 5.5 Hz) due to olefinic-H and the doublet at δ 2.08 (3H, J = 5.3 Hz) due to -CH₃ attached with the olefinic group were observed. The relative population of the *anti* (more stable *E* configuration) and the *syn* (less stable Z configuration) diastereomers can also be determined from the relative intensities of two quartets at δ 7.56 and 7.11, and it comes out E:Z = 2.6:1. These observations clearly indicate the presence of two diastereomeric 2,4-dinitrophenylhydrazone derivatives of acetaldehyde. This conclusively proves the generation of acetaldehyde in the reaction mixture as the reaction product. The presence of ammonia in the reaction product was tested by Nessler's reagent. The product CO₂ was qualitatively detected by bubbling nitrogen gas through the reaction mixture and passing the liberated gas through limewater. The stoichiometry of the reaction involving one of the gold(III) species may thus be written as

$$AuCl_{4}^{-} + {}^{+}H_{3}N(CH_{3})CHCOO^{-} + H_{2}O \rightarrow$$
$$AuCl_{2}^{-} + CH_{3}CHO + CO_{2} + NH_{4}^{+} + H^{+} + 2Cl^{-} (1)$$

Polymerization Test

Participation of free radicals, if any, was examined by the polymerization of initially added acrylonitrile (20% v/v) in the reaction mixture. Neither any haziness nor any precipitate appeared even after 3 h, thereby indicating the absence of involvement of oneelectron transfer process in the reduction of Au(III) to Au(I).

RESULTS

Effect of Reactant Concentrations

The oxidation of alanine by gold(III) was studied at different initial [Au^{III}] ((1.82-4.25) × 10^{-3} mol dm⁻³), keeping [alanine], [Cl⁻], pH, and temperature constant at 4.0×10^{-2} mol dm⁻³, 5.0×10^{-2} mol dm⁻³, 4.45, and 303 K, respectively. In all the kinetic experiments, [alanine] and [Cl⁻] were present in a large excess compared to [Au^{III}]. Acetate buffer was used to keep the [H⁺] at a constant value. The kinetic experiments were thus followed under pseudo-first-order conditions. The value of k_{obs} was found to be $(10.06 \pm 0.50) \times 10^{-4}$ s^{-1} , thereby indicating that the rate is of first order with respect to Au(III). The influence of variation of the alanine concentration on the reaction rate was also studied at constant [Au(III)], [Cl⁻], and pH of 3.04×10^{-3} mol dm⁻³, 5.0×10^{-2} mol dm⁻³, and 4.45, respectively, and at four different temperatures. The rate was found to increase with an increase in the alanine concentration (Table I). At each temperature, the plot of $k_{\rm obs}$ vs. [alanine] was linear passing through the origin (Fig. 1), indicating a simple first-order dependence on the alanine concentration. The reaction thus appeared to follow the rate law:

$$k_{\rm obs} = -\frac{(1/[{\rm Au}^{\rm III}])d[{\rm Au}^{\rm III}]}{dt} = k_2'[{\rm Alanine}] \qquad (2)$$

where k'_2 is the second-order rate constant.

Table I Variation of Pseudo-First-Order Rate Constants with Alanine Concentrations at Different Temperatures

10^2 [Alanine] (mol dm ⁻³)	$10^4 k_{\rm obs} ({\rm s}^{-1})$				
	298 K	303 K	308 K	313 K	
1.0	2.07 (2.07)	2.54 (2.54)	3.25 (3.25)	4.02 (4.02)	
2.0	4.30 (2.15)	5.33 (2.67)	6.45 (3.23)	8.18 (4.09)	
3.0	6.49 (2.16)	7.68 (2.56)	9.65 (3.22)	12.3 (4.10)	
4.0	8.60 (2.15)	10.1 (2.52)	13.0 (3.26)	16.5 (4.13)	
6.0	13.4 (2.24)	15.2 (2.53)	19.6 (3.27)	24.4 (4.06)	
8.0	16.2 (2.03)	20.1 (2.51)	25.3 (3.16)	33.0 (4.13)	
10.0	22.07 (2.21)	26.0 (2.60)	32.5 (3.25)	40.8 (4.08)	

 $[Au(III)] = 3.04 \times 10^{-3} \text{ mol } dm^{-3}, [H^+] = 3.55 \times 10^{-5} \text{ mol } dm^{-3}, and [Cl^-] = 0.05 \text{ mol } dm^{-3}.$

Figures in the parentheses represent the values of the second-order rate constants ($k'_2 = k_{obs}/[alanine], dm^3 mol^{-1} s^{-1}$).

 k'_2 values are of the order of 10^{-2} .



Figure 1 Observed rate constants k_{obs} as a function of [alanine] at four different temperatures. [Au(III)] = 3.04×10^{-3} mol dm⁻³, [H⁺] = 3.55×10^{-5} mol dm⁻³, and [Cl⁻] = 0.05 mol dm⁻³.

Effect of the Hydrogen Ion Concentration

The effect of variation of $[H^+]$ on the pseudo-firstorder rate constant was studied by changing the pH of the buffer solution from 3.72 to 4.77, keeping [Au(III)], [alanine], [Cl⁻], and temperature constant (Table II). No attempt was made to keep the ionic strength constant since the values of k_{obs} remained unchanged at different ionic strengths varied by the addition of NaClO₄. The rate was found to be inhibited by an increase in $[H^+]$. At a constant $[Cl^-]$, the plot of k'_2 vs. $[H^+]^{-1}$ (Fig. 2) was initially linear but showed a curvature at higher $[H^+]^{-1}$. The relation between k'_2 and $[H^+]$ may be expressed as

$$k'_{2} = \frac{k_{\rm b} + k_{\rm a}[{\rm H}^{+}]^{-1}}{1 + k_{\rm c}[{\rm H}^{+}]^{-1}}$$
(3)

Table II Second-Order Rate Constants (k'_2) and Different Empirical Constants $(k_a, k_b, and k_c)$ Obtained from the Acid Effect

$[Cl^{-}] \pmod{dm^{-3}}$	pН	$10^2 k'_2 (\mathrm{dm^3 \ mol^{-1} \ s^{-1}})$	$10^7 k_a (s^{-1})$	$10^3 k_{\rm b} ({\rm dm}^3 {\rm mol}^{-1} {\rm s}^{-1})$	$10^5 k_{\rm c} ({\rm mol} {\rm dm}^{-3})$
0.05	3.72	1.30	14.4 ± 1.9	8.88 ± 0.68	5.00 ± 0.76
	3.85	1.46			
	3.94	1.55			
	4.05	1.62			
	4.27	1.90			
	4.45	2.15			
	4.77	2.50			
0.20	3.72	0.99	3.54 ± 0.56	8.76 ± 0.25	1.00 ± 0.29
	3.85	1.05			
	3.94	1.10			
	4.05	1.14			
	4.27	1.31			
	4.45	1.43			
	4.77	1.85			

 $[Au(III)] = 3.04 \times 10^{-3} \text{ mol } dm^{-3}, [alanine] = 4.0 \times 10^{-2} \text{ mol } dm^{-3}, and 298 \text{ K}.$



Figure 2 Variation of the second-order rate constant with $[H^+]$. Nonlinear plot of k'_2 vs. $[H^+]^{-1}$ at 298 K. $[Au(III)] = 3.04 \times 10^{-3} \text{ mol dm}^{-3}$, $[alanine] = 4.0 \times 10^{-2} \text{ mol dm}^{-3}$, and $[Cl^-] = 0.05 \text{ mol dm}^{-3}$.

where k_a , k_b , and k_c are empirical constants. The curvefitting procedure has been employed to find out the empirical constants. The values of k_a , k_b , and k_c at two different [Cl⁻] are presented in Table II. It is evident that k_a and k_c are dependent on [Cl⁻], whereas k_b is independent of [Cl⁻].

Effect of the Chloride Ion Concentration

The influence of varying $[Cl^-]$ (0.05–0.4 mol dm⁻³) on the pseudo-first-order rate constant was studied by

the addition of NaCl, keeping [Au(III)], [alanine], and temperature constant at two different pH values, 4.05 and 4.45 (Table III). Chloride ion was found to have an inhibiting effect on the reaction rate. The plot of k'_2 vs. $[Cl^-]^{-1}$ was again found to be of similar nature as in the case of [H⁺] variation (Fig. 3). Thus at a constant [H⁺], the dependence of the second-order rate constant (k'_2) on $[Cl^-]$ may be expressed as

$$k'_{2} = \frac{k_{\rm e} + k_{\rm d} [\rm Cl^{-}]^{-1}}{1 + k_{\rm f} [\rm Cl^{-}]^{-1}}$$
(4)

pН	$10 [Cl^{-}] (mol dm^{-3})$	$10^2 k'_2 (\text{dm}^3 \text{ mol}^{-1} \text{ s}^{-1})$	$10^3 k_{\rm d} ({\rm s}^{-1})$	$10^3 k_{\rm e} ({\rm dm}^3 {\rm mol}^{-1} {\rm s}^{-1})$	$10^2 k_{\rm f} \ ({\rm mol} \ {\rm dm}^{-3})$
4.45	0.5	2.12	1.77 ± 0.34	9.92 ± 0.53	5.73 ± 1.5
	1.0	1.74			
	1.5	1.60			
	2.0	1.44			
	2.5	1.39			
	3.0	1.35			
	4.0	1.24			
4.05	0.5	1.69	0.74 ± 0.07	9.50 ± 0.11	2.20 ± 0.4
	1.0	1.38			
	1.5	1.26			
	2.0	1.20			
	2.5	1.14			
	3.0	1.12			
	4.0	1.07			

Table IIISecond-Order Rate Constants and Different Empirical Constants Obtained from the Cl⁻ Effect

 $[Au(III)] = 3.04 \times 10^{-3} \text{ mol dm}^{-3}, [alanine] = 4.0 \times 10^{-2} \text{ mol dm}^{-3}, \text{ and } 298 \text{ K}.$



Figure 3 Variation of the second-order rate constant with [Cl⁻]. Nonlinear plot of k'_2 vs. [Cl⁻]⁻¹ at 298 K. [Au(III)] = 3.04×10^{-3} mol dm⁻³, [alanine] = 4.0×10^{-2} mol dm⁻³, and pH = 4.45.

where k_d and k_f are [H⁺]-dependent constants, whereas k_e is independent of [H⁺] (Table III). It has been observed that at comparatively low pH and higher [Cl⁻] the plot of k'_2 vs. [Cl⁻]⁻¹ becomes linear (Fig. 4). It is plausible that at lower pH and [Cl⁻] $\gg k_f$ the rate law

(Eq. (4)) simplifies to

$$k'_{2} = k_{\rm e} + k_{\rm d} [{\rm Cl}^{-}]^{-1}$$
(5)

This limiting form of the rate law is also applicable to the initial linear part of the plot (Fig. 3) where $[Cl^-]$ is



Figure 4 Variation of the second-order rate constant with [Cl⁻]. Linear plot of k'_2 vs. [Cl⁻]⁻¹ at 298 K. [Au(III)] = 3.04×10^{-3} mol dm⁻³, [alanine] = 4.0×10^{-2} mol dm⁻³, and pH = 3.72.

comparatively high. It has been observed that the H⁺ independent constant, k_e , is almost identical at all the three studied pH values (3.72, 4.05, and 4.45).

Effect of Solvent

The effect of solvent composition on the rate of the reaction was studied by varying 1,4-dioxane (0–25% v/v), keeping [Au(III)], [alanine], pH, [Cl⁻], and temperature constant at 3.04×10^{-3} mol dm⁻³, 4.0×10^{-2} mol dm⁻³, 4.45, 0.05 mol dm⁻³, and 305 K, respectively. The reaction rate was found to be retarded by the addition of dioxane. Thus, the rate constant decreases with the decreasing dielectric constant of the medium.

Effect of Temperature

Variation of pseudo-first-order rate constants for the oxidation of alanine by Au(III) was studied at different [alanine] at four different temperatures. The values are presented in Table I. The linear plot of log (k'_2/T) vs. T^{-1} enables one to evaluate the enthalpy of activation (ΔH^{\neq}) using Eq. (6), followed by the determination of the entropy of activation (ΔS^{\neq}) .

$$\log(k_2'/T) = [\log(k_{\rm B}/h) + \Delta S^{\neq}/2.303 R] - \Delta H^{\neq}/2.303 RT$$
(6)

The values of ΔH^{\neq} and ΔS^{\neq} are found to be (31 ± 2) kJ mol⁻¹ and $-(174 \pm 7)$ J K⁻¹ mol⁻¹, respectively.

DISCUSSION

It is well known that in an aqueous solution the following equilibria between the cationic, zwitterionic, and anionic forms of alanine exist:

$$CH_{3}CH(\overset{\top}{\mathbf{N}}H_{3})COOH$$

$$\xrightarrow{K_{1}} CH_{3}CH(\overset{+}{\mathbf{N}}H_{3})COO^{-} + H^{+}$$
(7)

$$\begin{array}{c} \text{CH}_{3}\text{CH}(\text{NH}_{3})\text{COO}^{-} \\ \xrightarrow{K_{2}} \\ \hline \end{array} \text{CH}_{3}\text{CH}(\text{NH}_{2})\text{COO}^{-} + \text{H}^{+} \end{array} (8)$$

The values of pK_1 and pK_2 are 2.348 and 9.866, respectively, at 298 K [40]. Since the experiments were carried out in the pH range 3.72–4.77, it is quite evident that the concentration of the anionic species is negligible compared to the other two species of alanine. Among the zwitterionic and the cationic species, the former is expected to be the effective reducing agent since it is present in a large excess over the latter

and also contains the attacking nucleophile carboxylate anion [34]. The oxidant tetrachloroauric(III) acid is known to involve the following equilibria in an aqueous solution [41,42]:

$$HAuCl_4 \xrightarrow{K_3} H^+ + AuCl_4^-$$
(9)

$$\operatorname{AuCl}_{4}^{-} + \operatorname{H}_{2}\operatorname{O} \xleftarrow{K_{4}} \operatorname{AuCl}_{3}(\operatorname{OH}_{2}) + \operatorname{Cl}^{-} (10)$$

$$\operatorname{AuCl}_{3}(\operatorname{OH}_{2}) \xleftarrow{K_{5}} \operatorname{AuCl}_{3}(\operatorname{OH}^{-}) + \operatorname{H}^{+}$$
(11)

where the values of the equilibrium constants K_3 , K_4 , and K_5 are 1.0, 9.5×10^{-6} , and 0.25, respectively, at 298 K. Under the conditions of the experiment (pH 3.72–4.77 and [Cl⁻]: 0.05–0.30 mol dm⁻³), using the values of K_3 , K_4 , and K_5 , it can be shown that both HAuCl₄ and AuCl₃(OH₂) will be present in a negligible concentration. Hence, it becomes evident that the predominant and effective oxidizing species of gold(III) in the present system are AuCl₄⁻ and AuCl₃(OH)⁻, as observed in our recent communication [34]. Therefore, it may be proposed that the reaction takes place via two parallel paths involving the zwitterions and the oxidizing species, AuCl₄⁻ and AuCl₃(OH)⁻, as follows:

$$AuCl_{4}^{-} + {}^{+}H_{3}NCH(CH_{3})COO^{-} \xrightarrow{k_{1}}{slow} CH_{3}CH=NH_{2}^{+}$$
$$+CO_{2} + AuCl_{2}^{-} + H^{+} + 2Cl^{-}$$
(12)
$$AuCl_{3}(OH)^{-} + {}^{+}H_{3}NCH(CH_{3})COO^{-} \xrightarrow{k_{1}}{slow}$$
$$CH_{3}CH=NH_{2}^{+} + CO_{2} + AuCl_{2}^{-} + H_{2}O + Cl^{-}$$

This mechanism is in agreement with the oxidation of glycine [34] and other amino acids [14]. It may be suggested that an unstable intermediate is formed in a slow step between the zwitterion and gold(III) species

(13)

suggested that an unstable interinediate is formed in a slow step between the zwitterion and gold(III) species followed by its decomposition to produce the iminic cation (Eq. (14)).



$$\overset{H}{CH_{3}}C = \overset{+}{NH_{2}} + CO_{2} + AuCl_{2} + H^{+} + Cl^{-}$$
(14)

In several previous studies, the involvement of similar short-lived gold(III) intermediate complex has been reported in the oxidation of amino acids [34, 43] as well as other organic substrates [33]. In the final step, the iminic cation hydrolyzes rapidly to give acetaldehyde and NH_4^+ [44].

$$CH_3CH = \stackrel{+}{N}H_2 \xrightarrow[fast]{H_2O} CH_3CHO + NH_4^+$$
(15)

The involvement of similar iminic cation intermediate and its hydrolysis to give the corresponding aldehyde has been proposed in the oxidation of amino acids by Mn(VII) [14], Bi(V) [9], and *N*-bromoacetamide [45]. In the oxidation of glycine [3], DL-alanine [46], β -phenylalanine [46, 47], and DL-leucine [48], the reported products were the respective aldehydes along with NH⁴₄ and CO₂.

Based on the proposed mechanism, the rate of consumption of Au(III) may be expressed as

$$-\frac{d[\operatorname{Au(III)}]}{dt} = \{k_1[\operatorname{AuCl}_4^-] + k_3[\operatorname{AuCl}_3(\operatorname{OH})^-]\}[\operatorname{Ala}]$$
(16)

If $C_0 = [Au(III)]$ and $x = [AuCl_3(OH^-)]$, Eq. (16) may be written as

$$-\frac{d[\operatorname{Au}(\operatorname{III})]}{dt} = \{k_1(C_0 - x) + k_3 x\}[\operatorname{Ala}]$$
(17)

Expressing the value of x in terms of K_4 and K_5 , the above Eq. (17) is transformed to

$$-\frac{1}{[Au(III)]} \frac{d[Au(III)]}{dt}$$
$$= k_{obs} = \frac{k_1[H^+][Cl^-] + k_3K_4K_5}{K_4K_5 + [H^+][Cl^-]} [Ala] \quad (18)$$

Therefore, the second-order rate constant, k'_2 , is

$$k_{2}' = \frac{k_{\text{obs}}}{[\text{Ala}]} = \frac{k_{1}[\text{H}^{+}][\text{Cl}^{-}] + k_{3}K_{4}K_{5}}{K_{4}K_{5} + [\text{H}^{+}][\text{Cl}^{-}]}$$
(19)

Equation (19) is equivalent to the empirical rate expressions (3)–(5) provided one assumes that $k_a = k_3 K_4 K_5/[\text{Cl}^-]$, $k_b = k_1$, $k_c = K_4 K_5/[\text{Cl}^-]$, $k_d = k_3 K_4 K_5/[\text{H}^+]$, $k_e = k_1$, and $k_f = K_4 K_5/[\text{H}^+]$. This equivalent nature suggests that the mechanism depicted in steps (12)–(15) is reasonable, at least mathematically. It is clearly evident from the above relations that k_a and k_c are [Cl⁻]-dependent constants, k_d and k_f are [H⁺]-dependent constants, whereas k_b and k_e are independent of both [H⁺] and [Cl⁻] and this is in agreement with the experimental observations

(Tables II and III). The values of the rate constants $(k_1 \text{ and } k_3)$ for the rate-determining steps (12) and (13), respectively, have been evaluated for different sets of experiments. The average values of k_1 and k_3 are found to be $(0.882 \pm 0.046) \times 10^{-2}$ and $(3.01 \pm 0.43) \times 10^{-2}$ dm³ mol⁻¹ s⁻¹ obtained from acid effect, whereas respective values obtained from the chloride effect are $(0.97 \pm 0.024) \times 10^{-2}$ and $(2.79 \pm 0.29) \times 10^{-2}$ dm³ mol⁻¹ s⁻¹. The values are in good agreement with each other. The value of k_3 has been found to be higher than that of k_1 , which indicates the greater reactivity of $AuCl_3(OH)^-$ over $AuCl_4^-$. This fact agrees well with earlier reports [31, 34, 36]. Owing to nonavailability of the values of dissociation constants (K_3, K_4, K_5) at 303, 308, and 313 K, the values of k_1 and k_3 could not be determined at these temperatures. Consequently, the calculated values of ΔH^{\neq} and ΔS^{\neq} correspond to the activation parameters for the overall reaction. The reaction system involves several equilibria as well as the rate-determining steps and hence is very much complicated. Therefore, it is very difficult to state clearly which step plays the key role to contribute to the high negative value of ΔS^{\neq} . However, the rate-limiting step may have significant contribution to this ΔS^{\neq} . Here the transition state in the rate-limiting step becomes highly solvated due to polarization developed in it (as it is going to break down into a number of oppositely charged ions). This is associated with an immobilization of an appreciable number of solvent molecules (i.e., increase in electrostriction) [49], which results in a loss of entropy.

Considering the equilibrium steps (9–11) involving the different Au(III) species, it is evident that an increase in either $[H^+]$ or $[Cl^-]$ has an adverse effect on the concentration of the most reactive species of Au(III), namely, AuCl₃(OH)⁻. Consequently, the rate of the reaction diminishes with increasing $[H^+]$ or $[Cl^-]$. The rate has been found to decrease with an increase in the proportion of dioxane in the reaction medium (Table IV). A logical explanation is that with

Table IVEffect of Solvent on the Pseudo-First-OrderRate Constant at 305 K

Dioxane (%, v/v)	ε	$10^4 k_{\rm obs} ({\rm s}^{-1})$	
0	78.6	12.3	
5	74.2	9.98	
10	69.7	7.98	
15	65.3	6.91	
20	60.8	6.58	
25	56.4	5.76	

 $[Au(III)] = 3.04 \times 10^{-3} \text{ mol } dm^{-3}, [alanine] = 4.0 \times 10^{-2} \text{ mol } dm^{-3}, \\ [H^+] = 3.55 \times 10^{-5} \text{ mol } dm^{-3}, \text{ and } [Cl^-] = 0.05 \text{ mol } dm^{-3}.$

a decrease in the dielectric constant of the medium, the electrostatic attraction between the oppositely charged species increases and consequently the concentration of the effective oxidizing Au(III) species as well as the reactive zwitterionic form of alanine diminishes, thereby decreasing the reaction rate. It is interesting to note that the intermediate cyclic complex in the ratelimiting step decomposes to a number of oppositely charged ions, and hence it is quite expected that the transition state will be very much polar. Consequently, a decrease in the dielectric constant of the medium will disfavor the transition state and hence will decrease the reaction rate.

The reaction mechanism in this study has been found to be very much similar to that for the Au(III)glycine reaction. However, the reaction rate is slower than that for the oxidation of glycine. This may be accounted for by the fact that introduction of a $-CH_3$ group to the α -carbon of glycine increases the steric crowding in the transition state, and this is possibly reflected in the higher value of overall enthalpy of activation of Au(III)–alanine reaction over that of Au(III)– glycine reaction.

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