Three-Electron Oxidations. IX. Chromic Acid Oxidation of Glycolic Acid[1,2]

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Abstract: The chromic acid oxidation of glycolic acid follows the rate law $-d[Cr(VI)]/dt = K_1[HCrO_4^-][S](k_2 + k_3[S])$, where [S] is the concentration of glycolic acid and [Cr(VI)] the total analytical concentration of chromium(VI); $K_1 = 2.2$, $k_2 = 3.0 \times 10^{-3}\ M^{-1}\ sec^{-1}$, $k_3 = 1.3 \times 10^{-3}\ M^{-2}\ sec^{-1}$. The first term corresponds to a reaction leading to the formation of glyoxylic acid, formaldehyde and carbon dioxide in a 2:1:1 ratio. Glyoxylic acid is formed by two-electron oxidation with chromium(VI) and chromium(V) while chromium(IV) is responsible for the oxidative cleavage to formaldehyde and a $\cdot CO_2^-$ radical. The second term of the rate law corresponds to a three-electron oxidation in which a 2:1 glycolic acid–chromic acid complex decomposes directly to chromium(III), glyoxylic acid, and a $HO\dot{C}HCO_2H$ radical. No cleavage to formaldehyde takes place under conditions under which the second term of the rate law is dominant. The free radicals formed in either step react with chromium(VI) to yield an observable chromium(V) intermediate. The chromium(V) oxidation of glycolic acid yields glyoxylic acid only. The kinetic isotope effect in the oxidation of $HOCD_2CO_2H$ is $k_H/k_D = 6.15$ for k_2 and $k_H/k_D > 36.5$ for k_3. The high value of the isotope effect of the rate constant k_3 for the second term provided convincing proof that two carbon–hydrogen bonds are broken simultaneously and that the reduction of chromium(VI) to chromium(III) does take place in a single three-electron step.

Our recent study of the chromic acid oxidation of oxalic acid has shown that this reaction can best be described in terms of a mechanism involving a three-electron oxidation taking place within a complex formed from two molecules of the substrate and one molecule of chromic acid.[3] It is of considerable interest to determine whether this mechanism is unique for oxalic acid or whether it represents an example of a more general type of mechanism.

The study reported in this paper deals with the oxidation of glycolic acid. We felt that α-hydroxy acids are structurally sufficiently related to oxalic acid to offer a suitable class of compounds for further investigations. Moreover, we have recently shown that α-hydroxy acids greatly accelerate the rate of the reaction between chromic acid and isopropyl alcohol,[4] in a manner strongly resembling a similar effect observed earlier for oxalic acid[5] and offered strong evidence suggesting that the cooxidation of isopropyl alcohol and of α-hydroxy acids also proceeds by a three-electron mechanism.[6]

The chromic acid oxidation of α-hydroxy acids has been the subject of a large number of investigations particularly during the past 10 years,[7-17] although the oldest investigation was carried out more than 100 years ago.[18] The general conclusion was that the reaction resembles in many respects the oxidation of alcohols and it was therefore believed that it follows essentially the same mechanism. However, several differences were noted: (a) products of C–C cleavage were obtained,[12] (b) the reaction appears to be accelerated by the presence of the carboxyl group despite the electronegative character of this group which should have reduced the activity of the alcoholic group toward oxidation,[9] (c) kinetic evidence of the formation of an intermediate complex between a hydroxy acid and chromic acid was obtained.[15] No satisfactory mechanism which could explain all the above features has been advanced. No indication that more than one molecule of hydroxy acid may be involved in the transition state of the oxidation has been recorded.

In this paper, we wish to report the results obtained in a detailed investigation of the chromic acid oxidation of glycolic acid which we selected as the simplest example of an α-hydroxy acid with an oxidizable hydroxyl group.

Experimental Section

Materials. Glycolic acid (Eastman Kodak) was crystallized from petroleum ether (mp 74–76°). Sodium dichromate (J. T. Baker, Reagent) and diglycolic acid (Eastman Kodak) were used without further purification. Ethyl glycolate (Eastman Kodak) was purified by vacuum distillation and the purity of the compound verified by GLPC. Perchloric acid solutions were prepared from 60% perchloric acid (B & A Reagent).

$HOCD_2CO_2H$. Lithium aluminum deuteride (4 g) dissolved in ether (50 ml) was added dropwise to a solution of diethyl oxalate (75 ml) in ether (75 ml). The reaction mixture was heated to reflux for 1 hr and after cooling slowly added to a 10% solution of sulfuric acid containing crushed ice. Ethyl glycolate-d_2 was extracted twice from the reaction mixture with 100-ml portions of ether. The ether extract also contained some ethylene glycol-d_4 and unreacted ethyl oxalate. The extract was further extracted once with 50 ml of water which left the ethyl oxalate in the organic layer and ethyl glycolate-d_2 and ethylene glycol-d_4 in the aqueous layer. Ethyl glycolate-d_2 was separated from the glycol by extraction with 100 ml of ether, the former being much more soluble in ether than the latter. The ether extract was checked with GLPC which showed a single peak for ethyl glycolate. The extract was dried over anhydrous magnesium sulfate and the ether was removed by distillation. On hydrolysis of the ethyl glycolate-d_2 with aqueous KOH, the potassium salt of glycolic-d_2 acid precipitated out. Glycolic-d_2 acid was obtained by acidification with 1 M perchloric acid. The yield after two crystallizations from petroleum ether was 1.5 g, mp 76–78°. The NMR spectrum of the compound showed the absence of peaks due to CH_2 protons.

Kinetic Measurements. Reaction rates were determined spectrophotometrically at 350 nm on Cary 14, Cary 15, and Zeiss PMQII spectrophotometers equipped with thermostated cell holders. The initial chromium(VI) concentration in all kinetic runs was $6.28 \times 10^{-4}\ M$. For glycolic acid, pseudo-first-order rate constants were calculated from the slopes of the linear parts of the log (absorbance) vs. time plots (Figure 1); rate constants obtained from multiple determinations were within ±5% of each other.

Diglycolic acid gave good first-order plots. For ethyl glycolate the solution of chromic acid and perchloric acid in the cell was allowed to reach the required temperature in the thermostated cell holder and the reaction was started by injecting the required amount of ethyl glycolate from a syringe. Rate constants were determined from the initial part of the plots (Figure 2). The rate acceleration observed in the course of the reaction was due to hydrolysis of the ester. (When ethyl glycolate was first dissolved in aqueous perchloric acid, kept in a thermostat for about 20 min, and a solution of chromic acid added afterwards, much higher reaction rates were obtained since an almost complete hydrolysis of the ester had taken place.)

Product Analysis. (A) **Glycolic Acid Oxidation.** In a typical experiment glycolic acid (18.7 g), perchloric acid (6.5 ml, 9.69 *M*) and sodium dichromate (1.0 ml, 1.745 *M*) were mixed together in a 100-ml volumetric flask and the total volume was made up to the mark with distilled water. After completion of the reaction, 25 ml of the reaction mixture was extracted about 15 times, because of the low solubility of oxalic acid in ether,[19] with 10-ml portions of ether. An experiment with a known quantity (~1 mg) of oxalic acid in 25 ml of water showed that extraction with 100 ml of ether (ten times with 10-ml portions) can remove at least 95% of it. The extract was dried over anhydrous magnesium sulfate and the ether removed under vacuum. The residue was tested for formic acid[20] by the chromotropic acid test[21] and for oxalic acid by the Aniline Blue test.[21]

The yield of formic acid present in the crude reaction mixture was determined spectrophotometrically.[22] The relative yields of formic and oxalic acids were determined by gas chromatography using a 6-ft diethylene glycol succinate column at 100°. The method was tested by injecting a mixture containing known quantities of oxalic, formic, and glycolic acids. The other possible product, glyoxylic acid, has a much higher boiling point and is not expected to interfere.

To the remaining 50 ml of the reaction mixture a solution of 2,4-dinitrophenylhydrazine in 2 *N* HCl was added and the precipitated 2,4-dinitrophenylhydrazone was filtered, washed, dried, and weighed, and was identified as essentially pure glyoxylic acid 2,4-dinitrophenylhydrazone by melting point (194–195°) and mixture melting point (193–195°) and by comparison of the NMR spectrum (in acetone-d_6) with that of an authentic sample.

In oxidation experiments where the glycolic acid:chromium(VI) ratio was low, the NMR spectrum of the 2,4-dinitrophenylhydrazone indicated the presence of formaldehyde (δ 6.8–7.7) along the glyoxylic acid. The ratios of the two products were determined from integrated NMR spectra. This procedure was tested and found to give satisfactory results with synthetic mixtures of the two dinitrophenylhydrazones.

At very high glycolic acid concentrations (over 4 *M*) the 2,4-dinitrophenylhydrazone of glyoxylic acid did not precipitate. The dinitrophenylhydrazone formed in this case was determined spectrophotometrically in a methanol–KOH solution at 420 nm (ϵ = 4.02×10^4).[5,23] The extinction coefficient was determined by measuring the absorbance at 420 nm for a series of solutions of the dinitrophenylhydrazone of known concentrations. Beers' law was found to be valid in the range under observation. Under these conditions a formaldehyde yield as low as 5% can be safely detected.

Carbon dioxide was determined manometrically in a Warburg apparatus.[3] The gas evolved was found to be only carbon dioxide since no gas evolution was observed when sodium hydroxide was placed in the "well" of the Warburg flask during reaction. This excludes the possibility of the formation of carbon monoxide.

Product yields were calculated from the expressions yield (%) = 100 $nP/3(\mathrm{Cr(VI)})$, where Cr(VI) and *P* are millimoles of chromium(VI) consumed and of millimoles of a given product formed; *n* has the value of 2 for glyoxylic acid and formaldehyde, and 4 for formic and oxalic acid. For carbon dioxide the yield was calculated for CO_2 formed in excess over the amount formed from the oxidative cleavage of glycolic and glyoxylic acid from the expression

$$\mathrm{CO_2\ yield\ (\%)} = \frac{100}{(\mathrm{Cr(VI)})}[\mathrm{CO_2} - (\mathrm{CH_2O} + \mathrm{HCO_2H})]$$

(B) **Glyoxylic Acid Oxidation.** In a typical experiment glyoxylic acid (3.0 ml, 0.652 *M*), perchloric acid (3.2 ml, 9.69 *M*), distilled water (13.6 ml), and sodium dichromate (0.2 ml, 1.745 *M*) were allowed to react at 25°. After the completion of the reaction, 2,4-dinitrophenylhydrazine in 2 *N* HCl was added to 10 ml of the reaction mixture, and the dinitrophenylhydrazone was filtered, washed, and dried. The NMR spectrum of the precipitate in acetone-d_6 showed the absence of peaks due to formaldehyde. The remaining 10 ml of the reaction mixture was tested for formic acid and oxalic acid by the method described before and positive tests were obtained for both acids.

(C) **Test for Free Radicals.** In a typical experiment glycolic acid (1.0 ml, 3.6 *M*), distilled water (3.9 ml) and a saturated solution of mercuric chloride (5.0 ml) were mixed in a 50-ml erlenmeyer flask. Sodium dichromate (0.1 ml, 1.745 *M*) was added to the mixture. A precipitate of mercurous chloride appeared as the reaction proceeded, indicating the formation of free radicals during the reaction.

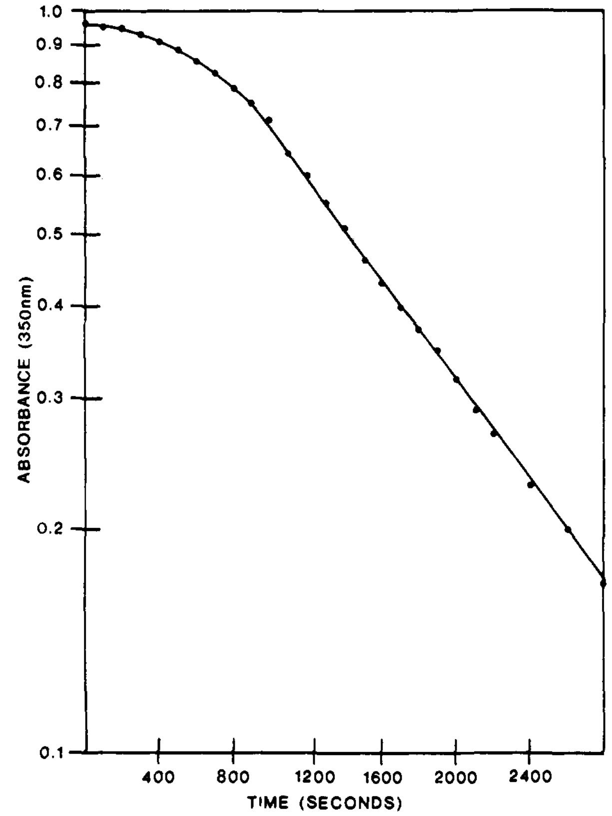


Figure 1. First-order rate plot for chromic acid oxidation of glycolic acid at 25°: glycolic acid = 0.546 *M*, perchloric acid = 0.485 *M*, initial chromium(VI) = 6.28×10^{-4} *M*; 1-cm cell.

Results and Discussion

Products. Table I lists the products obtained from the chromic acid oxidation of glycolic acid at different glycolic acid and chromium concentrations.

As long as a sufficient excess of glycolic acid over chromic acid is maintained, glyoxylic acid, formaldehyde, and carbon dioxide are the only oxidation products. At high concentration of glycolic acid, glyoxylic acid is formed in quantitative yields. As the concentration of glycolic acid is decreased, the formation of formaldehyde and carbon dioxide in a strictly 1:1 ratio is observed. At the lowest concentration of glycolic acid, the yield of glyoxylic acid and formaldehyde is 67 and 33%, respectively.

When a lesser excess of glycolic acid over chromic acid is employed, larger amounts of carbon dioxide are formed. In addition, formic and oxalic acids appear among the oxidation products. The fact that these products are observed only at higher chromic acid:glycolic acid ratios clearly indicates that they are formed by further oxidation of glyoxylic acid. An experiment in which glyoxylic acid was subjected to chromic acid oxidation under similar conditions indeed yielded formic acid, oxalic acid, and carbon dioxide.

Rate Studies. Figure 1 shows a typical example of a log (absorbance) vs. time plot. After an initially curved portion, a good straight line plot is obtained. The linear part of the plot was used to determine the values of the experimental pseudo-first-order rate constants.

The initial curvature is due to the formation of a chromium(V) intermediate. A similar effect was observed earlier in the chromic acid oxidation of oxalic acid.[3] Figure 3 shows the time dependence of the absorption at 750 nm where chromium(V) is the only absorbing species.[24] The

Table I. Products of the Chromic Acid Oxidation of Glycolic Acid (Perchloric Acid = 0.628 *M*)

Glycolic acid,[a] *M*	Chromium-(VI),[a] 10^3 *M*	Glyoxylic acid 10^5 mol	Glyoxylic acid Yield, %	Formaldehyde 10^5 mol	Formaldehyde Yield, %	Carbon dioxide 10^5 mol	Carbon dioxide Yield, %	Oxalic acid 10^4 mol	Oxalic acid Yield, %	Formic acid 10^4 mol	Formic acid Yield, %	Total
6.00[b]	0.455[b]	0.68	98.9									98.9
2.73	0.455	5.91	87.0	1.02	14.7	0.99	14.5					101.7
1.64	0.455	5.44	80.0	1.25	18.4	1.10	16.1					98.4
0.498	0.455	5.00	73.6	1.70	25.0	1.60	23.4					98.6
		4.91[c]	72.3[c]	1.75[c]	25.8[c]							98.1
0.273	0.455	4.60	67.7	1.90	27.9	1.98	29.0					95.6
0.110	0.455	4.60	67.7	2.26	33.3	2.20	32.2					101.0
0.055	0.455	4.60	67.7	2.20	32.7	2.25	33.0					100.4
6.80[b]	4.92[b]	7.45	101.0									101.0
2.46	35.0	146	27.8			190.0	35.7	3.3	12.5	6.50	24.8	100.8
1.30	70.0	242	2.3	1.01	0.96	480.0	54.1	12.0	22.9	10.0	19.0	99.3
0.39	35.0	100	1.9	1.88	3.6	280.0	70.3	3.2	12.3	3.17	12.1	100.2
0.23	70.0	34.0	0.33	12.5	11.9	720.0	85.0					97.2
0.071	35.0			83.4	15.9	390.0	88.6					104.5

[a] Total volume = 100 ml. [b] Total volume = 10 ml. [c] With 0.114 *M* acrylonitrile.

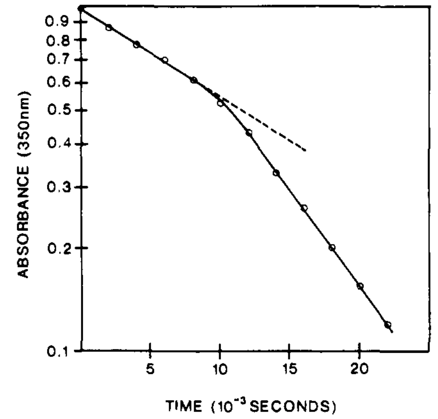


Figure 2. Chromic acid oxidation of ethyl glycolate at 25°: ethyl glycolate = 0.216 *M*, perchloric acid = 0.628 *M*, initial chromium(VI) = 6.28×10^{-4} *M*.

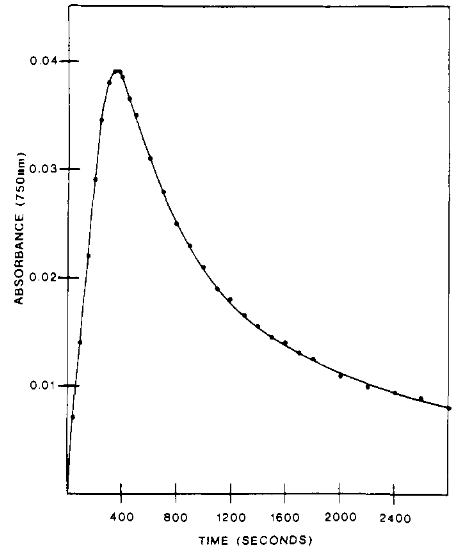


Figure 3. Time dependence of absorbance of 750 nm at 25°: glycolic acid = 0.498 *M*, perchloric acid = 0.620 *M*, initial chromium(VI) = 6.28×10^{-3} *M*; 5-cm cell.

formation of a chromium(V) intermediate is clearly observable.

The dependence of the experimental first-order rate constant on the concentration of glycolic acid at four different acidities is given in Figure 4. Also given are initial rate constants for ethyl glycolate and rate constants for diglycolic acid (Table II) at a single acidity. The latter two com-

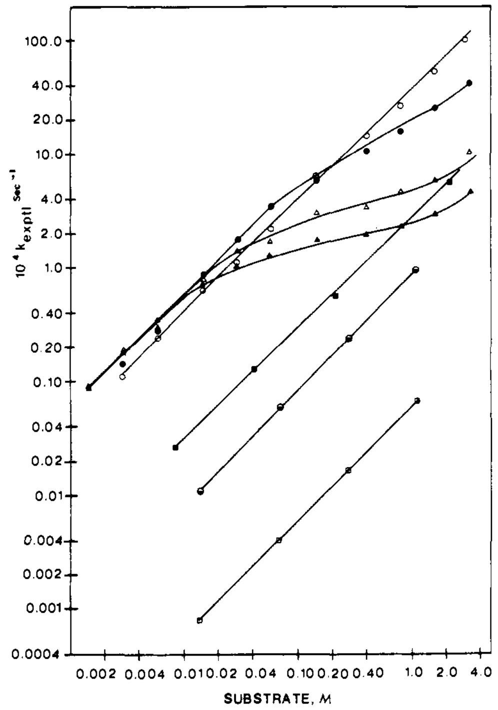


Figure 4. Effect of substrate concentration on rates of chromic acid oxidations: chromium(VI) = 6.28×10^{-4} *M*; glycolic acid, 25° (○) $HClO_4$ = 1.26 *M*, (●) $HClO_4$ = 0.628 *M*, (△) $HClO_4$ = 0.126 *M*, (▲) $HClO_4$ = 0.063 *M* (the lines are drawn through the points calculated from eq 11); ethyl glycolate, 25° (■) $HClO_4$ = 0.628 *M*; diglycolic acid (◒) 60°, $HClO_4$ = 0.628 *M*, (□) 25° (estimated from rates at 60°).

Table II. Chromic Acid Oxidation of Ethyl Glycolate and Diglycolic Acid (Perchloric Acid = 0.628 *M*, Initial Chromium(VI) = 6.28×10^{-4} *M*)

Substrate, *M*	*T*, °C	$10^5 k_{\text{exptl}}$, sec^{-1}	$10^4 k_{\text{exptl}}/[S]$
		Ethyl Glycolate	
0.0086	25	0.259	3.01
0.043	25	1.30	3.02
0.216	25	5.75	2.66
2.16	25	59.2	2.74
		Diglycolic Acid	
0.0137	60	0.118	0.860
0.0685	60	0.608	0.888
0.274	60	2.41	0.879
1.096	60	9.63	0.878
1.096	25	0.679	0.062

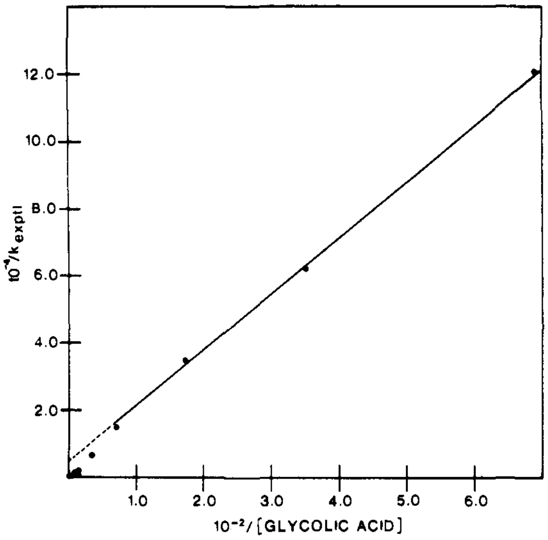


Figure 5. Determination of equilibrium constant K_1 for chromic acid oxidation of glycolic acid at 25° (perchloric acid = 0.628 M).

pounds show first-order dependency over the whole range and are considerably less reactive than glycolic acid[15] with diglycolic acid (which lacks a free hydroxyl group) exhibiting a particularly strongly reduced reactivity.[26]

The oxidation of glycolic acid is much more complicated. At low concentrations of glycolic acid the reaction is first order in glycolic acid. At lower perchloric acid concentrations reaction rates are acidity independent, while high acidity reduces the rates noticeably.

At higher glycolic acid concentrations and high acidity (1.26 M $HClO_4$) the reaction remains first order in glycolic acid almost over the whole range of glycolic acid concentrations (1.5×10^{-3} to 3.2 M) with only a small deviation observable at very high glycolic acid concentrations. At low acidities and intermediate glycolic acid concentrations the order in glycolic acid decreases and appears at first to approach a zero-order dependence, a behavior which is characteristic of reactions in which large concentrations of an intermediate complex of reactant and substrate are formed. At still higher concentrations of glycolic acid, however, the experimental rate constant starts increasing again (instead of reaching a constant value as would have been expected if only a single intermediate were formed).

In order to describe the observed kinetic behavior we had to postulate a reaction scheme in which an intermediate can either decompose directly or react with a second substrate molecule (Scheme I)

Scheme I

$$\text{Cr(VI)} + \text{S} \xrightleftharpoons{K'_1} \text{C}_1 \qquad (1)$$

$$\text{C}_1 \xrightarrow{k'_2} \text{products} \qquad (2)$$

$$\text{C}_1 + \text{S} \xrightarrow{k'_3} \text{products} \qquad (3)$$

where Cr(VI) = $HCrO_4^-$ + H_2CrO_4, S represents the substrate (glycolic acid), and C_1 a 1:1 glycolic acid-chromic acid complex. This scheme leads to eq 4 (for constant acidity).

$$k_{\text{exptl}} = \frac{K'_1[\text{S}](k'_2 + k'_3[\text{S}])}{1 + K'_1[\text{S}]} \qquad (4)$$

From visual inspection of the k_{exptl} vs. [S] plot (Figure 4), it is obvious that the term K'_1[S] in the denominator will become important much earlier than the term k'_3[S]; otherwise a shift from first to second order in glycolic acid rather than a shift toward zero order would have been observed. At intermediate glycolic acid concentrations, eq 5 and 6 will therefore very approximately describe the dependence of the experimental rate constant on the concentration of the glycolic acid.

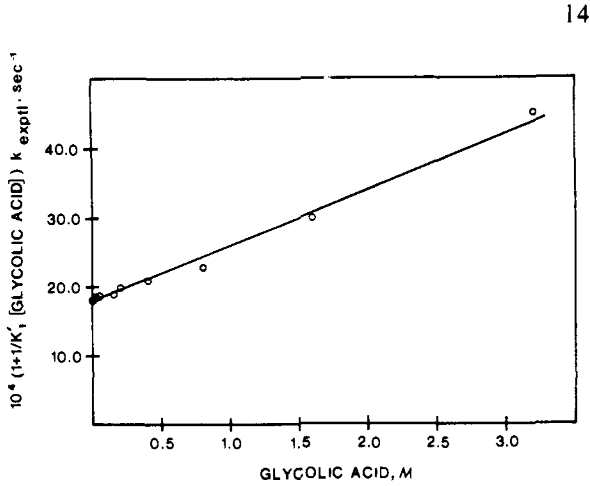


Figure 6. Determination of k'_3 for chromic acid oxidation of glycolic acid at 25° (perchloric acid = 0.628 M).

$$k_{\text{exptl}} = \frac{K'_1k'_2[\text{S}]}{1 + K'_1[\text{S}]} \qquad (5)$$

$$\frac{1}{k_{\text{exptl}}} = \frac{1}{K'_1k'_2[\text{S}]} + \frac{1}{k'_2} \qquad (6)$$

An example of a plot of $1/k_{\text{exptl}}$ vs. 1/[S] is given in Figure 5. It can be seen that the plot is indeed fairly linear over most of the range with the deviation from linearity occurring, as is to be expected, at high concentrations of glycolic acid.

Rearranging eq 4 gives eq 7 which permits the determination of k'_3 from the slope of the plot of the left hand side of the equation vs. the concentration of glycolic acid (Figure 6).

$$\left(1 + \frac{1}{K'_1[\text{S}]}\right)k_{\text{exptl}} = k'_2 + k'_3[\text{S}] \qquad (7)$$

The decrease in the oxidation rate at very high acidities and low concentrations of glycolic acid suggests that only the monochromate ion $HCrO_4^-$ and not the total concentration of chromic acid should be considered. A similar situation was found earlier in the chromic acid oxidation of oxalic acid.[3] Therefore, the dissociation of chromic acid (eq 8) was taken into account, and the experimental rate constants obtained for the measurements in 1.26 M $HClO_4$ were divided by $K_a/(K_a + h_0)$ before rate and equilibrium constants were calculated. The value of $h_0 = 2.51$ was used.[28]

$$H_2CrO_4 \xrightleftharpoons{K_a} H^+ + HCrO_4^-;\; K_a = 4.2^{27} \qquad (8)$$

The rate and equilibrium constants determined in this way are summarized in Table III. All three constants are

Table III. The Rate and Equilibrium Constants for Chromic Acid Oxidation of Glycolic Acid at 25°

$[H^+]$, M	K_1'	$K_1 = K_1'[H^+]$	$10^4k_2'$	$10^3k_2 = k_2'/[H^+]$	$10^4k_3'$	$10^3k_3 = k_3'/[H^+]$
2.51 (h_0)	1.09	2.74[a]	70	2.8	32	1.3
0.628	3.22	2.02	18	2.9	8	1.3
0.126	18.1	2.28	4.0	3.2	1.6	1.3
0.063	36.0	2.27	2.0	3.2	0.80	1.3
Average	$K_1 = 2.2$		$k_2 = 3.0 \times 10^{-3}$ M^{-1} sec^{-1}		$k_3 = 1.3 \times 10^{-3}$ M^{-2} sec^{-1}	

[a] Because of the large error involved in the determination of K_1' at high acidity, this value was omitted in the computation of the average.

Table IV. Chromic Acid Oxidation of Glycolic Acid at 25°

Glycolic acid, *M*	10^4k (sec^{-1}), Perchloric acid, *M*: 1.26 (h_0 = 2.51) Obsd	1.26 Calcd	0.628 Obsd	0.628 Calcd	0.126 Obsd	0.126 Calcd	0.063 Obsd	0.063 Calcd
3.20	99	115	42	41	10.2	8.9	4.8	4.5
1.60	53	63	25	29	5.8	6.5	3.2	3.3
0.80	26	33	16.1	20	4.8	5.1	2.5	2.6
0.40	14.4	16.9	10.6	13.7	3.5	4.2	2.2	2.3
0.145	6.4	5.9	6.0	6.7	3.0	2.9	1.78	1.67
0.058	2.2	2.36	3.4	3.3	1.68	1.94	1.30	1.29
0.029	1.10	1.16	1.77	1.75	1.41	1.29	1.02	0.96
0.0145	0.63	0.59	0.83	0.90	0.81	0.76	0.61	0.63
0.0058	0.24	0.24	0.28	0.38	0.36	0.35	0.24	0.32
0.0029	0.112	0.118	0.140	0.158	0.193	0.180	0.178	0.170
0.00145			0.085	0.079	0.088	0.090	0.086	0.085

acidity dependent. While the equilibrium constant K'_1 is inversely proportional to the acidity, both rate constants are directly proportional to concentration of hydrogen ions. Good acidity independent constants were obtained by multiplying or dividing respectively the equilibrium and rate constants by the first power of the concentration of hydrogen ions (using h_0 instead of $[H^+]$ at high acidities).

Rate Law. The full rate law for the chromic acid oxidation of glycolic acid is given by eq 9 where $[Cr_T]$ stands for

$$-\frac{d[Cr(VI)]}{dt} = k_{exptl}[Cr_T] = K_1[HCrO_4^-][S](k_2 + k_3[S]) \qquad (9)$$

the total analytical concentration of chromium(VI) (eq 10).

$$[Cr_T] = [Cr(VI)] + [C_1] = [HCrO_4^-] + [H_2CrO_4] + [C_1] \qquad (10)$$

The dependence of the experimental pseudo-first-order rate constants on the concentration of glycolic acid, chromium(VI), and acidity obtained from eq 9 and 10 is given in eq 11.

$$k_{exptl} = \frac{K_1[S](k_2 + k_3[S])}{1 + K_1[S]/[H^+] + [H^+]/K_a} \qquad (11)$$

Equation 11 was used to calculate values which are compared with the experimentally obtained data in Table IV and show an excellent agreement over a very wide range of concentrations. The equation was also used to construct calculated curves in Figure 4.

The rate law shows that the reaction proceeds by two routes: through a transition state containing one molecule of glycolic acid per one molecule of chromium(VI) and through a second one with two molecules of glycolic acid per molecule of chromic acid.

Figure 7 shows that the two kinetic terms can be correlated with the composition of the oxidation products. The yield of formaldehyde (and of carbon dioxide) is proportional to the fraction of the reaction which proceeds through the 1:1 complex. Under conditions under which the reaction proceeds essentially completely through the 1:1 complex, a 33% yield of formaldehyde and of carbon dioxide is obtained.

Mechanism. A mechanism which will account for all experimental data is given in Scheme II.

The acidity dependence of the observed equilibrium constants K'_1 shows that the formation of the complex between the acid chromate ion, $HCrO_4^-$ and glycolic acid proceeds with a loss of a proton. Consequently, the complex must be doubly negatively charged. This is a surprising observation insofar as in this respect the complex between chromic acid

Scheme II

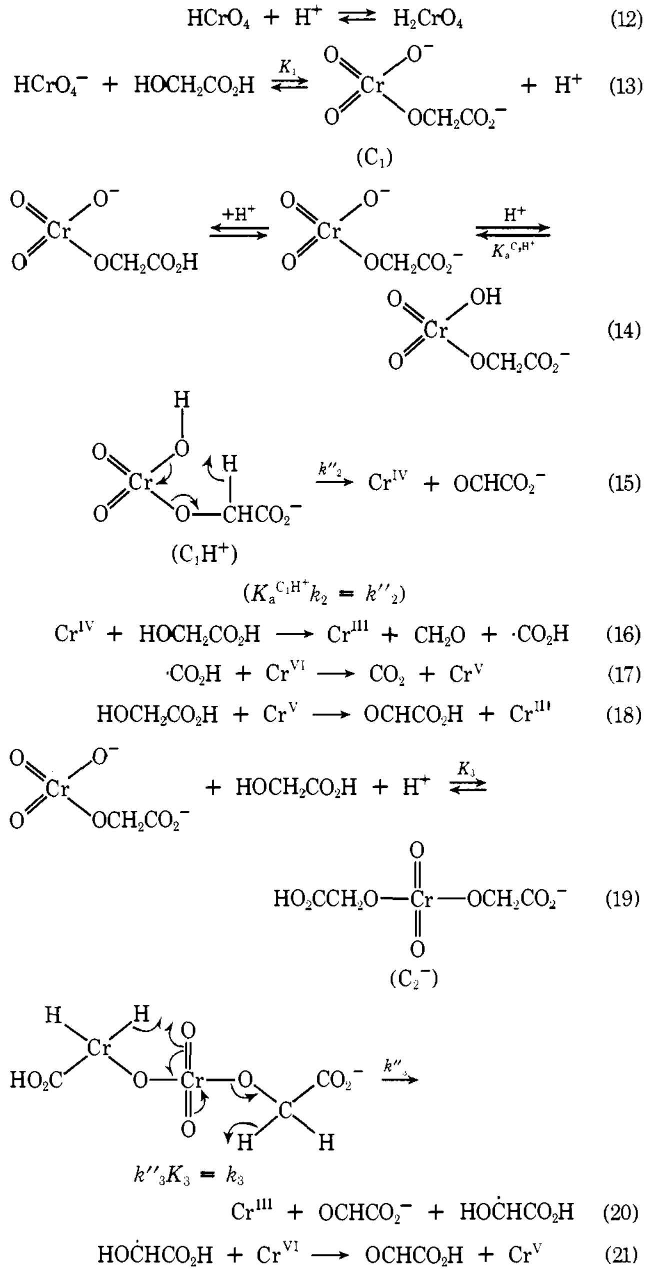


Table V. Chromic Acid Oxidation of Glycolic Acid-d_2 at 25° (Perchloric Acid = 0.628 *M*)

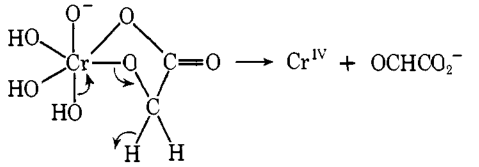
Glycolic acid, *M*	10^4k_H, sec^{-1}	10^4k_D, sec^{-1}	k_H/k_D
0.0145	0.826	0.134	6.15
3.20	42.0	3.34	12.6
6.01	133	3.65	36.5

and glycolic acid differs from the neutral complex formed between chromic acid and oxalic acid.

Kinetic data alone do not permit determining the actual structure of the complex; however, some conclusions can be drawn from the observation that both ethyl glycolate and diglycolic acid are considerably less reactive and that neither of them shows any deviation from first-order dependence which would indicate the formation of an intermediate. Obviously, both a free hydroxyl and carboxyl group are essential for enhanced reactivity and for the formation of an intermediate in large enough concentrations to be kinetically detectable.

The simplest structure of an intermediate is shown in Scheme II; other structures, e.g., a cyclic structure with hexacoordinated chromium and negative charges on Cr–O groups, would also be consistent with our data.[29]

The rate law shows that the oxidative decomposition of the 1:1 complex requires an additional proton. There is very little doubt that the carboxylate group in the chromate ester of glycolic acid will be considerably more basic than the CrO_3^- group and therefore protonated to a much larger extent. However, as oxidation reactions are generally facilitated by increasing the negative charge on the substrate and by the protonation of the oxidant, we believe that the chromic acid ester of the glycolate anion will be the actual intermediate in the oxidative decomposition if the oxidation involves an acyclic complex. Decomposition via an octahedral transition state could also be envisioned (eq 22).



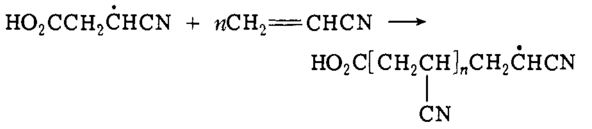
(22)

Chromium(IV) has been known for its ability to react with carbon-carbon bond cleavage in preference to carbon-hydrogen bond cleavage provided that the reaction leads to a relatively stable free radical or to the release of internal strain.[31] The 33% yield of formaldehyde and of carbon dioxide at low glycolic acid concentrations and the observation that the formation of carbon dioxide, but not of formaldehyde, is suppressed in the presence of acrylonitrile (Table I) demonstrate that *chromium(IV) oxidizes glycolic acid with carbon-carbon bond cleavage* and gives formaldehyde and a $\cdot CO_2H$ (or $\cdot CO_2^-$) free radical.

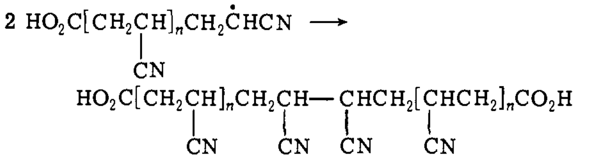
The radical then reacts with another molecule of chromium(VI) leading to the formation of a chromium(V) intermediate, the formation and decay of which can be observed at 750 nm (Figure 3). Chromium(V) reacts with glycolic acid to yield another molecule of glyoxylic acid.

In the presence of acrylonitrile the following reactions take place (eq 23-25)[31c]

$$\cdot CO_2H + CH_2{=}CHCN \longrightarrow HO_2CCH_2\dot{C}HCN \quad (23)$$



$$HO_2CCH_2\dot{C}HCN + nCH_2{=}CHCN \longrightarrow HO_2C[CH_2\underset{\displaystyle CN}{\underset{|}{C}}H]_nCH_2\dot{C}HCN \quad (24)$$



$$2\ HO_2C[CH_2\underset{\displaystyle CN}{\underset{|}{C}}H]_nCH_2\dot{C}HCN \longrightarrow HO_2C[CH_2\underset{\displaystyle CN}{\underset{|}{C}}H]_nCH_2\underset{\displaystyle CN}{\underset{|}{C}}H{-}\underset{\displaystyle CN}{\underset{|}{C}}HCH_2[\underset{\displaystyle CN}{\underset{|}{C}}HCH_2]_nCO_2H \quad (25)$$

(or disproportionation)

The second kinetic term in the rate law (eq 9) corresponds to a transition state containing two molecules of glycolic acid. Although we have no direct evidence that a discrete 2:1 intermediate of glycolic acid and chromic acid is formed, we believe that the formation of such an intermediate represents the most plausible interpretation. The rate law indicates that the transition state should be negatively charged. In the absence of any information on the equilibrium reaction between the 1:1 complex and glycolic acid, we assume that the 2:1 complex will also carry one negative charge. The fact that we were unable to detect its formation kinetically indicates that the equilibrium constant governing its formation must be quite low. As the second term in the rate law is important and the oxidation via the 2:1 complex becomes the main reaction path at high glycolic acid concentrations, the rate of the oxidative decomposition of this complex must be high.

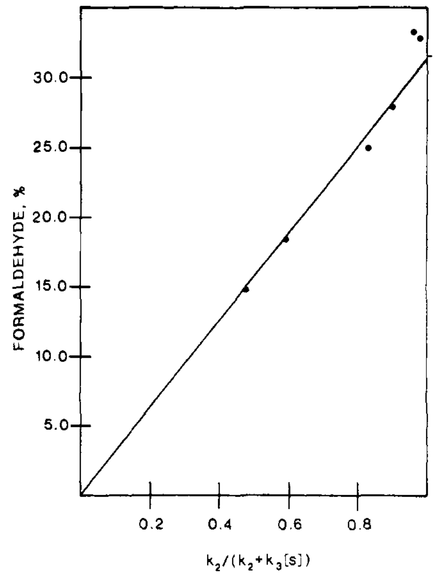
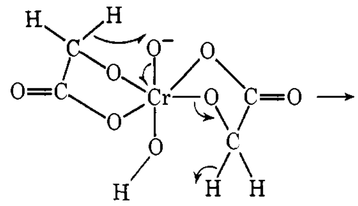


Figure 7. Dependence of the yield of formaldehyde on the fraction of glycolic acid oxidized through the 1:1 complex.

We have earlier suggested that complexes of chromic acid with two molecules of substrates owe their high reactivity to their ability of reacting by a one-step three-electron oxidation mechanism in which chromium(VI) is reduced directly to chromium(III), by-passing thus the formation of the very unstable chromium(IV) intermediate. In the present case, additional support for this interpretation is supplied by the observation that the reaction products formed through the 2:1 complex contain no formaldehyde or carbon dioxide. As the chromium(IV) oxidation of glycolic acid leads to carbon-carbon bond cleavage to formaldehyde and $\cdot CO_2H$, the absence of cleavage products provides conclusive evidence that no chromium(IV) intermediate is formed in the oxidative decomposition of the 2:1 complex; consequently, this complex must decompose directly to chromium(III).

In order to account for the sole formation of glyoxylic acid we assume that one molecule of glyoxylic acid and a free radical $HO\dot{C}HCO_2H$ is formed in this step; thus carbon-hydrogen bond cleavage must take place both in the two- and one-electron oxidations of the substrate. The next step of the reaction must be again the oxidation of the intermediate free radical by chromium(VI) leading to glyoxylic acid and chromium(V) which will react as shown above (eq 18).

At this point we do not have enough evidence which would permit us to make specific conclusions about the structure of the 2:1 complex between glycolic acid and chromic acid. The structure given in the scheme simply represents one of several plausible possibilities; an octahedral transition state (eq 26) could also be suggested.



$$\mathrm{Cr^{III}} + \mathrm{OCHCO_2^-} + \mathrm{HOCHCO_2H} \quad (26)$$

Isotope Effects. Table V gives the observed deuterium isotope effect with glycolic acid-d_2, at both high and low glycolic acid concentrations.

At the lowest concentration, where the reaction is first order in the substrate ($k_2(\mathrm{S}) \gg k_3[\mathrm{S}]^2$) and the oxidation goes almost entirely (>99.4%) through the 1:1 complex, the isotope effect is very close to the normal value predicted for the breaking of one carbon-hydrogen bond.[32] At high glycolic acid concentrations, where oxidation through the 2:1 complex becomes important, large isotope effects are observed; the value obtained for 6 M glycolic acid is of an unprecedented magnitude.[33] If one made the assumption that rate and equilibrium constants determined at relatively dilute solutions are still valid at high concentrations then one could estimate the actual value for the isotope effect of the k_3 term from eq 27 obtained by rearranging eq 11. The

$$k_3^{\mathrm{D}} = \frac{k_{\mathrm{exptl}}(1 + K_1[\mathrm{S}]/[\mathrm{H}^+] + [\mathrm{H}^+]/K_a)}{K_1[\mathrm{S}]^2} - \frac{k_2^{\mathrm{D}}}{[\mathrm{S}]} \quad (27)$$

value of k_2^{D} ($2.99 \times 10^{-3}/6.15 = 0.486 \times 10^{-3}$) was obtained from k_2 (Table III) and $k_{\mathrm{H}}/k_{\mathrm{D}}$ for low glycolic acid concentrations (Table V). When this equation is applied to the value of $k_{\mathrm{exptl}} = 3.34 \times 10^{-4}$ sec^{-1} obtained for the oxidation of glycolic acid-d_2 (3.20 M), a value of $k_3^{\mathrm{D}} = 3.1 \times 10^{-5}$ corresponding to an isotope effect of $k_3^{\mathrm{H}}/k_3^{\mathrm{D}} \approx 41$ is obtained.

However, eq 11 does not hold well for very concentrated solutions of glycolic acid. This becomes apparent when it is used to calculate the rate constant for the 6.0 M solution; the calculated value (63×10^{-4} sec^{-1}) is only about one-half of the experimental value (133×10^{-4} sec^{-1}). The higher value of the observed rate constant suggests that a considerably larger proportion of the reaction proceeds via the 2:1 complex than would be estimated from equilibrium constants determined in dilute solutions. The same factors are probably responsible for eq 27 giving a too high value for k_3^{D} leading to an unreasonably high ($k_{\mathrm{H}}/k_{\mathrm{D}} \sim 60$) isotope effect for k_3. It is therefore more realistic to assume that in the very concentrated 6.01 M solution the bulk of the oxidation proceeds via the 2:1 complex and that the observed isotope effect in this solution approaches the correct value for $k_3^{\mathrm{H}}/k_3^{\mathrm{D}}$.

The magnitude of the isotope effect observed for the glycolic acid oxidation at high substrate concentrations shows clearly that *two carbon-hydrogen bonds are broken simultaneously in the rate limitation decomposition of the 2:1 complex. This finding and the similar observation in the cooxidation of glycolic acid and isopropyl alcohol*[1] *provide convincing evidence for the synchronous nature of three-electron oxidations.*

References and Notes

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