Contribution from the Department of Chemistry,
Jadavpur University, Calcutta-700032, India

# Kinetics of Oxidation of Glyoxylic and Pyruvic Acids by Vanadium(V)

KALYAN K. SEN GUPTA* and HRIDAY RANJAN CHATTERJEE

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The kinetics of oxidation of glyoxylic and pyruvic acids by vanadium(V) have been studied spectrophotometrically in the presence of perchloric acid over a wide range of temperature. Each reaction is first order with respect to both vanadium(V) and the substrates. The order with respect to hydrogen ion is significantly less than unity in each reaction. The empirical rate law for both reactions is $-\mathrm{d}[\mathrm{V(V)}]/\mathrm{d}t = k_sK[\mathrm{V(V)}][\text{substrate}][\mathrm{H^+}]/(1 + K[\mathrm{H^+}])$. The energies of activation have been evaluated as 18.4 ± 1.0 and 12.8 ± 0.5 kcal $\mathrm{mol^{-1}}$ for the respective reactions. The corresponding entropies of activation are 1.5 ± 3.0 and −17.0 ± 1.5 cal $\mathrm{mol^{-1}\ deg^{-1}}$. The reactions seem to proceed through the formation of free-radical intermediates.

## Introduction

$\alpha$-Keto acids occupy an important position in several metabolic sequences and are compounds of considerable biological interest. Glycine degradation to $CO_2$ may occur via glyoxylic acid arising from transamination or from oxidative deamination of glycine under the influence of the specific enzyme glycine oxidase. Glyoxylic acid formation from glycine is one of the routes of glycine catabolism.[1a] Pyruvic acid is a compound of exceptional importance for the understanding of carbohydrate metabolism[1b] because it can undergo a wide variety of reactions and give rise to many different products. However, the mechanisms of oxidations of both glyoxylic and pyruvic acids are yet to be understood. The mechanism of chromic acid oxidation of organic compounds has been the subject of numerous studies in recent years.[2] The kinetics of the oxidation of some organic compounds by vanadium(V) have been studied.[3] No attempt seems to have been made to study the oxidation of $\alpha$-keto acids by vanadium(V) although the kinetics and mechanism of oxidation of glyoxylic and pyruvic acids by chromium(VI)[4a] have been reported. The oxidations of glyoxylic and pyruvic acids by vanadium(V) have been incorporated in this report.

## Experimental Section

**Reagents.** Inorganic materials were of either BDH (AnalaR) or E. Merck (GR) grades. Glyoxylic acid (Fluka, AG) and pyruvic acid (Sigma, biochemical grade) were used. A standard solution of vanadium(V) was prepared by dissolving ammonium metavanadate in standard perchloric acid. The strength of vanadium(V) was determined by mixing a known volume of the solution in an excess of standard Mohr's salt solution. The excess Mohr's salt was titrated against standard dichromate solution using diphenylamine sulfonate as indicator.

**Absorption Spectra.** The standard solution of vanadium(V) was reduced to vanadium(IV) by excess glyoxylic acid. Spectra of different vanadium(IV) solutions in the range $3.17 \times 10^{-3}$ to $2.54 \times 10^{-2}$ M were recorded at different wavelengths from 500 to 950 nm on a Beckman Model DU spectrophotometer using a cell of path length 1 cm. All solutions show absorption maxima at 765 nm. The molar extinction coefficient ($\epsilon$) for vanadium(IV) at this wavelength has been found to be 16.8 ± 0.2. Beer's law has been found to be valid in the concentration ranges of vanadium(IV) studied.

**Kinetic Measurements.** The rate of formation of vanadium(IV) was followed spectrophotometrically in the above-mentioned spectrophotometer by measuring the absorbance of the reaction mixture at 765 nm where vanadium(IV) shows an absorption maximum. A cell of path length 1 cm was used throughout. The cell compartment was kept at constant temperature by circulating water through its thermospacer coils from a thermostat. The reaction was initiated by adding a requisite quantity of the substrate into the solution of the oxidant and perchloric acid, all maintained at constant temperature. Mixing time was taken as zero time. Equivalent amounts of the substrate and oxidant were used in all runs except in the case of finding the effect of substrate where $[\text{vanadium(V)}]_0 \ll [\text{substrate}]_0$. In the former cases, the reductions followed second-order kinetics whereas in the latter case the reductions followed first-order kinetics. The second-order kinetics were followed within 8% for 2 half-times (see plot 1 of Figure 1). Generally six to eight experimental points were noted in each run. Duplicate measurements were reproducible to ±3%. Most of the runs were carried out in 1.0 M $HClO_4$ unless otherwise mentioned.

**Polymerization Test.** A series of reactions between vanadium(V) and each of glyoxylic acid and pyruvic acid were carried out in the presence of 20% (w/v) of acrylamide. The initial concentrations of glyoxylic and pyruvic acids were $9.7 \times 10^{-3}$ and $7.45 \times 10^{-3}$ M, respectively, whereas the vanadium(V) concentration was adjusted to $1.45 \times 10^{-2}$ M in each reaction. A viscous solution was noticeable at $[\mathrm{H^+}] \leq 0.1$ M in each case. The viscous solution obtained in the oxidation of glyoxylic acid turned to a thick gel when kept for 24 h. Blank experiments from which either vanadium(V) or $\alpha$-keto acids were excluded gave no detectable polymer during this time. The experiments do indicate that the reaction between vanadium(V) and $\alpha$-keto acids produce species which are capable of the initiation of acrylamide polymerization.

## Results

**Stoichiometry and Product Analyses.** Glyoxylic or pyruvic acid was mixed with a large excess of vanadium(V) in perchloric acid, and the mixture was kept for 24 h. Vanadium(IV) was then estimated spectrophotometrically. In the oxidations of glyoxylic and pyruvic acids, 1 mol of the substrate consumed 2 mol of vanadium(V). The reaction mixtures were shaken with ether, and the solvent was removed from the extract. The oxidation product formic acid in the oxidation of glyoxylic acid was tested[4b] by the addition of chromotropic acid which gave a violet color. Acetic acid obtained in the other reaction was tested[4b] by the addition of lanthanum nitrates and iodine to give a deep blue precipitate.

The oxidation product was analyzed by a conductometric method.[4c] Two inflections were obtained when the mixture was titrated by standard NaOH. The first inflection was due to the neutralization of highly ionized $HClO_4$, and the second one was due to the weak acid.[4c] It was found that 1 mol of each substrate produces 1 mol of organic acid in each reaction. The reactions may therefore be represented stoichiometrically by eq 1 and 2.

$$\mathrm{CHOCOOH + 2V(V) + H_2O \rightarrow HCOOH + 2V(IV) + CO_2 + 2H^+} \quad (1)$$

$$\mathrm{CH_3COCOOH + 2V(V) + H_2O \rightarrow CH_3COOH + 2V(IV) + CO_2 + 2H^+} \quad (2)$$

**Rate Law.** The rate of the reaction may be represented as eq 3, where $a$ and $b$ are the initial molar concentrations of the

$$\mathrm{d}x/\mathrm{d}t = k_2(a - x/2)(b - x) \quad (3)$$

substrate and the oxidant, respectively, and $x$ is the amount of vanadium(V) which has been reduced at any time $t$. When $b = 2a$, eq 3 changes to (4). Equation 4 on integration gives

$$\mathrm{d}x/\mathrm{d}t = {}^1/_2k_2(b - x)^2 \quad (4)$$

eq 5. Since $x \propto A_t$ and $b \propto A_\infty$, the plot of $1/(A_\infty - A_t)$

$$1/(b - x) = {}^1/_2k_2t + 1/b \quad (5)$$

against $t$ gives a straight line. Typical second-order plots at different temperatures have been shown in Figure 1.

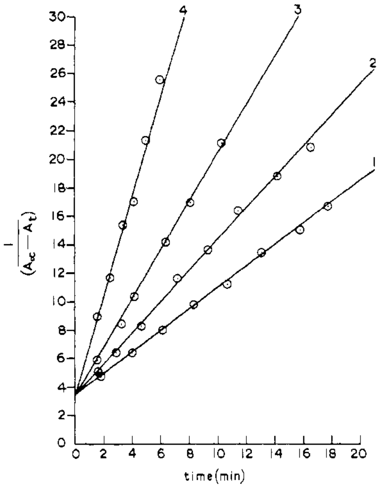


**Figure 1.** Second-order plots at different temperatures. Plots of $1/(A_\infty - A_t)$ against time for the oxidation of glyoxylic acid by vanadium(V): (1) $T$ = 28.0 °C; (2) $T$ = 31.5 °C; (3) $T$ = 36.0 °C; (4) $T$ = 43.0 °C. [Glyoxylic acid] = $9.3 \times 10^{-3}$ M; [vanadium(V)] = $1.86 \times 10^{-2}$ M; [$HClO_4$] = 1.0 M.

**Table I.** Effect of [substrate] on Pseudo-First-Order Rate Constants[a]

| | | | | | | |
|---|---|---|---|---|---|---|
| $10^2$[glyoxylic], M | 0.5 | 1.0 | 2.0 | 3.0 | 4.0 | 5.0 |
| $10^3k_{obsd}$, $s^{-1}$ | 0.16 | 0.31 | 0.63 | 0.93 | 1.28 | 1.65 |
| $10^2k_{obsd}$/[glyoxylic], $M^{-1}$ $s^{-1}$ | 3.2 | 3.1 | 3.15 | 3.1 | 3.2 | 3.3 |
| $10^2$[pyruvic], M | 0.5 | 1.0 | 2.0 | 3.0 | 4.0 | 5.0 |
| $10^3k_{obsd}$, $s^{-1}$ | 0.466 | 0.92 | 1.83 | 2.80 | 3.76 | 4.58 |
| $10^2k_{obsd}$/[pyruvic], $M^{-1}$ $s^{-1}$ | 9.3 | 9.2 | 9.15 | 9.3 | 9.4 | 9.2 |

[a] [vanadium(V)] = $2.5 \times 10^{-3}$ M, [$HClO_4$] = 1.0 M, $T$ = 5.0 °C. [glyoxylic] = concentration of glyoxylic acid; [pyruvic] = concentration of pyruvic acid.

**Effect of [Substrate].** The reactions were studied at constant [vanadium(V)]$_0$ but at different [substrate]$_0$. The acid concentration was kept fixed at each run. The pseudo-first-order rate constant increases (Table I) with the increase in [substrate]$_0$. The plots of $1/k_{obsd}$ against 1/[substrate]$_0$ are linear passing through the origin in each reaction. The average values of $k_{obsd}$/[substrate]$_0$ at 5 °C are $(3.16 \pm 0.08) \times 10^{-2}$ and $(9.25 \pm 0.15) \times 10^{-2}$ $M^{-1}$ $s^{-1}$ for glyoxylic and pyruvic acids, respectively. The results indicate that the order with respect to each substrate and hence vanadium(V) is unity.

**Effect of [$HClO_4$].** The effect of [$H^+$] on the second-order rate constant was studied at different [$HClO_4$]$_0$ but constant [reactants]$_0$. The ionic strength was held constant at 2.0 M by adding requisite amounts of $NaClO_4$. In an attempt to correlate the rate of oxidation with acid concentration, both hypotheses[5,6] for the mechanism of acid catalysis were tested. For both reactions, Zucker–Hammett plots, i.e., log $k_2$ against log [$HClO_4$] and log $k_2$ against $-H_0$ (where $H_0$ is the Hammett acidity function), are linear. The plots indicate that the reactions are acid catalyzed. The slopes of log $k_2$ against log [$HClO_4$] plots are 0.28 and 0.26 for the respective reactions. The plots of $1/k_2$ against 1/[$H^+$] are also linear having intercepts as in Figure 2.

**Effect of $NaClO_4$.** The effects of the addition of $NaClO_4$ in the presence of $HClO_4$ for both reactions were studied at constant concentrations of the respective substrate, oxidant, and perchloric acid. No salt effect was observed at [$NaClO_4$] ≤ 0.2 M although the rate constant has a tendency to increase at higher concentrations of $NaClO_4$. The results are recorded in Table II.

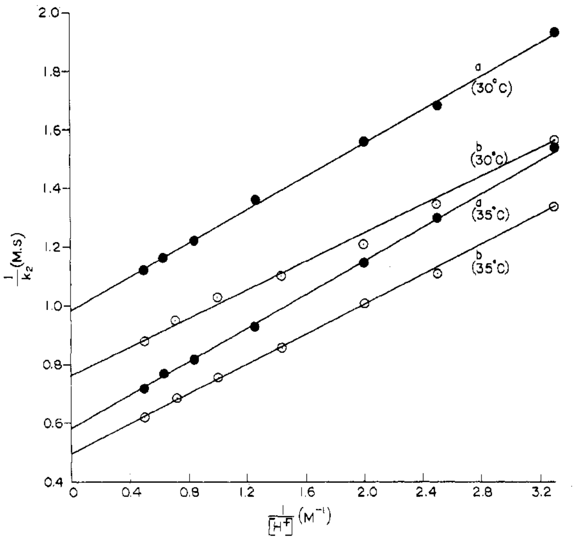


**Figure 2.** Plots of reciprocal second-order rate constants against the reciprocal of [$H^+$] for the oxidation of glyoxylic and pyruvic acids at $\mu$ = 2.0 M: (a) [glyoxylic acid] = $9.3 \times 10^{-3}$ M, [vanadium(V)] = $1.86 \times 10^{-2}$ M; (b) [pyruvic acid] = $4.43 \times 10^{-3}$ M, [vanadium(V)] = $8.87 \times 10^{-3}$ M.

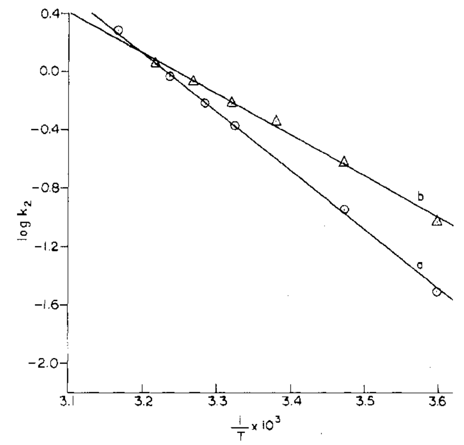


**Figure 3.** Influence of temperature on second-order rate constants and the plots of log $k_2$ against $1/T$: (a) [glyoxylic acid] = $9.30 \times 10^{-3}$ M, [vanadium(V)] = $1.86 \times 10^{-2}$ M, [$HClO_4$] = 1.0 M; (b) [pyruvic acid] = $7.45 \times 10^{-3}$ M, [vanadium(V)] = $1.49 \times 10^{-2}$ M, [$HClO_4$] = 1.0 M.

**Table II.** Effect of [$NaClO_4$] on Second-Order Rate Constants at [$HClO_4$] = 1.0 M

| | | | | |
|---|---|---|---|---|
| [$NaClO_4$],[a] M | 0.0 | 0.8 | 1.20 | 1.60 |
| $10K_2$,[a] $M^{-1}$ $s^{-1}$ | 5.7 | 6.7 | 7.2 | 7.5 |
| [$NaClO_4$],[b] M | 0.0 | 0.50 | 1.00 | 1.50 |
| $10K_2$,[b] $M^{-1}$ $s^{-1}$ | 5.8 | 7.2 | 8.4 | 9.6 |

[a] [glyoxylic] = $9.30 \times 10^{-3}$ M, [vanadium(V)] = $1.86 \times 10^{-2}$ M, $T$ = 30 °C. [b] [pyruvic] = $4.435 \times 10^{-3}$ M, [vanadium(V)] = $8.87 \times 10^{-3}$ M, $T$ = 28 °C.

**Effect of Temperature.** Second-order rate constants in 1 M $HClO_4$ at various temperatures were measured for both reactions. The plots of log $k_2$ against $1/T$ are linear (Figure 3) and from the gradients, the energies of activation ($E_a$) have been evaluated to be 18.4 ± 1.0 and 12.8 ± 0.5 kcal $mol^{-1}$ for the respective reactions. The corresponding values of the entropies of activation are 1.5 ± 3.0 and −17.0 ± 1.5 cal $mol^{-1}$ $deg^{-1}$. The values are not widely different from those obtained earlier[12] and suggest that free-radical mechanisms may also be operative in the present reactions. The details of the calculations of different thermodynamic parameters have been reported elsewhere.[7]

## Discussion

Jones and Waters[8a] have suggested that $VO_2^+$ and $VO(OH)^{2+}$ are the active species for the oxidation of $\alpha$-hydroxy acids. It was later shown[8b] that the hydrated form of vanadium(V), i.e., $V(OH)_4^+$ and not $VO_2^+$, remains in equilibrium with $V(OH)_3^{2+}$ according to eq 6. The observed hydrogen ion dependence on rate constant suggests that the oxidation is affected by both $V(OH)_4^+$ and $V(OH)_3^{2+}$. The steps of the reaction are

$$V(OH)_4^+ + H^+ \overset{K}{\rightleftharpoons} V(OH)_3^{2+} + H_2O \qquad (6a)$$

$$R + V(V) \xrightarrow{k_s} \text{products} \qquad (6b)$$

(R = hydrated α-keto acid)

The rate law would then be

$$\frac{-d[V(V)]}{dt} = \frac{k_s K[V(V)][R][H^+]}{1 + K[H^+]} \qquad (7)$$

Therefore

$$\frac{-d[V(V)]}{dt}\,\frac{1}{[V(V)][R]} = k_2 = \frac{k_s K[H^+]}{1 + K[H^+]} \qquad (8)$$

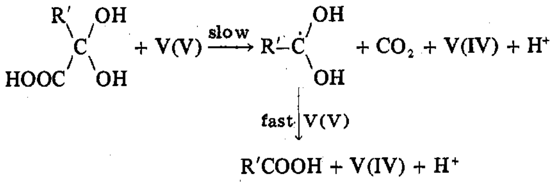
or

$$\frac{1}{k_2} = \frac{1}{k_s} + \frac{1}{k_s K[H^+]} \qquad (9)$$

A plot of $1/k_2$ against $1/[H^+]$ should give a straight line with a positive intercept at the $y$ axis. Such plots are shown in Figure 2. The equilibrium constants ($K$) have been calculated to be 3.44 and 3.40 at 30 °C whereas the same have been found to be 2.03 and 2.00 at 35 °C from the respective plots. All runs were taken at $\mu = 2.0$ M. The close agreement in $K$ values obtained by different reactions at each temperature substantiates the validity of rate equation 7. The heat of reaction ($\Delta H$) associated with equilibrium 6a has been evaluated to be $-19.6$ kcal mol$^{-1}$, using the integrated form of Vant Hoff's isochore equation.

It has been shown earlier[9,16] that most reactions involving vanadium(V) proceed via a free-radical mechanism wherein vanadium(V) undergoes a one-electron reduction. In the present investigations, the addition of acrylamide to the reaction mixture gave a polymeric product in the oxidations of both glyoxylic and pyruvic acids. This indicates that vanadium(V) behaves as a 1-equiv oxidant in the present investigations. However, the dissociation constants ($K_a$) of glyoxylic and pyruvic acids are $4.7 \times 10^{-4}$ and $3.2 \times 10^{-3}$ at 25 °C. It is also known that α-keto acids are hydrated[10,11] in aqueous solution and hence the hydrated forms of the respective substrates will remain as undissociated molecules since $[H^+] \gg K_a$ in the presence of high concentration of perchloric acid (~1.0 M). It is, therefore, suggested that vanadium(V) reacts with hydrated α-keto acid (R′ = H for glyoxylic acid and R′ = $CH_3$ for pyruvic acid) in the slow step to give a free radical, vanadium(IV), and $CO_2$. The free radical further picks up vanadium(V) and by a fast step gives the products of oxidation (see Scheme I).

**Scheme I**



The initial oxidation of glyoxylic acid to give oxalic acid followed by the oxidation of oxalic acid to give $CO_2$ seems unlikely since the oxidation of oxalic acid by vanadium(V) is kinetically very complex.[13] The observed stoichiometry as well as formic acid, which is obtained as one of the products of oxidation, further supports the mechanism suggested in the oxidation of glyoxylic acid. Again, it has been shown that acetaldehyde on oxidation[14,15] by metal ion in mineral acid gives both formic and acetic acids. The formation of formic acid from acetaldehyde is possible only when acetaldehyde enolizes. In the present investigation, however, initial decarboxylation of pyruvic acid followed by the oxidation of acetaldehyde seems unlikely since in that case two different products, e.g., formic and acetic acid, would result. The presence of only one product of oxidation, e.g., acetic acid, justifies the mechanism suggested in the oxidation of pyruvic acid.

**Registry No.** V(V), 22537-31-1; $H^+$, 12408-02-5; glyoxylic acid, 298-12-4; pyruvic acid, 127-17-3.

## References and Notes

(1) (a) A. White, "Principles of Biochemistry", 2nd ed, McGraw-Hill, New York, N.Y., p 548, 1954; (b) E. S. G. Barron, "Trends in Physiology and Biochemistry", Academic Press, New York, N.Y., 1952, p 471.
(2) (a) M. Rahman and J. Rocek, *J. Am. Chem. Soc.*, **93**, 5462 (1971); (b) F. Hasan and J. Rocek, *ibid.*, **94**, 9073 (1972); (c) F. Hasan and J. Rocek, *ibid.*, **94**, 3181 (1972).
(3) (a) J. S. Littler and W. A. Waters, *J. Chem. Soc.*, 1299, 3014, 4046 (1959); (b) E. Pellizzetti, E. Mentasti, and G. Saini, *Gazz. Chim. Ital.*, **104** (9, 10) (1974); (c) T. L. Riechel and D. T. Sawyer, *Chem. Abstr.*, **83**, 87 268 (1975).
(4) (a) K. K. Sen Gupta and T. Sarkar, *Tetrahedron*, **31**, 123 (1975); (b) F. Feigl, "Spot Tests in Organic Analysis", Elsevier, London, 1956, pp 331, 342; (c) J. Rose, "Advanced Physico-chemical Experiments", Pitman & Sons Ltd., London, 1964, p 104.
(5) L. P. Hammett, "Physical Organic Chemistry", McGraw-Hill, New York, N.Y., 1940, Chapter IX.
(6) M. A. Paul and F. A. Long, *Chem. Rev.*, **57**, 13 (1957).
(7) K. K. Sen Gupta and J. K. Chakladar, *J. Chem. Soc., Dalton Trans.*, 222 (1974).
(8) (a) J. R. Jones and W. A. Waters, *J. Chem. Soc.*, 2772 (1960); (b) R. N. Mehrotra, *J. Chem. Soc. B*, 642 (1968).
(9) J. S. Littler, A. I. Mallet, and W. A. Waters, *J. Chem. Soc.*, 2761 (1961).
(10) Z. Strehlow, *Z. Electrochem.*, **66**, 392 (1962).
(11) W. J. Albery, R. P. Bell, and A. L. Powel, *Trans. Faraday Soc.*, **61**, 1194 (1965).
(12) B. B. Pal and K. K. Sen Gupta, *Inorg. Chem.*, **14**, 2268 (1975).
(13) J. R. Jones and W. A. Waters, *J. Chem. Soc.*, 4757 (1961).
(14) J. Shorter and C. N. Hinshelwood, *J. Chem. Soc.*, 3276 (1950).
(15) J. Shorter, *J. Chem. Soc.*, 3425 (1950).
(16) K. Kustin and D. L. Toppen, *Inorg. Chem.*, **12**, 1404 (1973).