Bull. Chem. Soc. Jpn., ***68***, *3325—3329(1995)*

Quinolinium Dichromate Oxidation of Bicyclic Alcohols

Irona Nongkynrih and Mahendra K. Mahanti*

Department of Chemistry, North-Eastern Hill University, Shillong 793003, India

(Received January 24, 1995)

Kinetic data on the rates of quinolinium dichromate oxidation of some bicyclic alcohols are reported and discussed with reference to the factors that influence the nature of the transition state. The mechanistic pathway has been visualized via the formation of the intermediate chromic ester which undergoes decomposition to give the product. A cyclic transition state is suggested; being a Hückel-type system ($4n+2$), this would be an allowed process.

The effects of conformation of secondary alcohols, in general, on their rates of oxidation by chromic acid have been reported in earlier investigations.[1,2] The electronic effects on the rates of oxidation of secondary alcohols under different reaction conditions have been rationalized.[3] It has been observed that the rates of oxidation of secondary alcohols with chromic acid have been linked to the extent of steric strain during the conversion of the alcohol to the ketone, that is, a change from the sp^3 to the sp^2 hybridized state.[4,5] In our sustained efforts to exploit quinolinium dichromate (QDC) for the oxidation of organic substrates[6] this paper presents the kinetic features of the oxidation of some bicyclic alcohols (borneol and isoborneol) by QDC in acidic medium, using dimethylformamide (DMF) as the solvent, under a nitrogen atmosphere. The rates of oxidation of borneol and isoborneol have been compared with those of norborneols (α- and β-) and cyclohexanol, and the effect of steric strain relief on the rates of decomposition of the chromic esters have been examined.

Experimental

Materials All the substrates were obtained from the Aldrich Chemical Co., and were purified by either recrystallization or distillation (cyclohexanol), and their purity was checked from physical constants. Isoborneol-2-*d* was prepared using a method reported earlier.[7] Quinolinium dichromate (QDC) was prepared by the reported method[8] and its purity was checked by estimating Cr(VI) iodometrically. *N*,*N*-dimethylformamide (BDH, Poole, UK) was distilled under reduced pressure before use. Sulfuric acid (E. Merck) was used as such after a check of its physical constants.

Kinetics. The reactions were done under pseudo-first-order conditions, maintaining a large excess of the substrate with respect to QDC. The reactions were done at constant temperature (±0.1 K), and were followed by monitoring the absorption band at 440 nm, spectrophotometrically (Systronics, digital model), as described earlier.[6] All reactions were done under a nitrogen atmosphere. The rate constants were evaluated from the linear ($r>0.996$) plots of log [QDC] against time. The values reported were the mean of two or more runs, and were reproducible to within ±3%. The solvent was DMF or DMF–water mixtures. The reaction mixtures remained homogeneous in the solvent systems used.

Product Analysis. Solutions of the substrate (0.01 M) and QDC (0.001 M) taken in DMF containing H_2SO_4 (1 M) were mixed and left for 24 h under a nitrogen atmosphere ($1\ M=1\ mol\,dm^{-3}$). Solvent ether was added to the reaction mixture, and the organic layer was washed with water, dried over anhydrous $MgSO_4$, and then concentrated. The product, obtained as colorless crystals (mp 179 °C), had a carbonyl band at $1733\ cm^{-1}$ and certain other IR bands characteristic of camphor. The 2,4-dinitrophenylhydrazone (DNP) of camphor was prepared and recrystallized from ethanol to give orange crystals (yield 85—90%; mp 175 °C).

Results and Discussion

The oxidation of borneol and isoborneol by QDC in DMF resulted in the formation of camphor. Under these experimental conditions, there was no further oxidation of the ketone.

Stoichiometry. The stoichiometric experiments were done under nitrogen at 323 K, under the conditions of $[QDC]_0 > [substrate]_0$, at varying acid concentrations. The disappearance of Cr(VI) was followed, spectrophotometrically, until the absorbance values became constant. The $[QDC]_\infty$ was estimated. Stoichiometric ratios, $\Delta[QDC]/\Delta[Substrate]$, of 0.33 and 0.35 for borneol and isoborneol, respectively, were obtained (Table 1), which conformed to the overall equation:

$$3C_{10}H_{18}O + 2Cr^{6+} \longrightarrow 3C_{10}H_{16}O + 2Cr^{3+} + 6H^+ \qquad (1)$$

Rate Law. Under pseudo-first-order conditions, the individual kinetic runs were first-order with respect to QDC. Further, the pseudo-first-order rate constant (k_1) did not alter appreciably with changing QDC concentrations (10-fold range), at constant substrate concentration (large excess). this indicated a first-order dependence on QDC (Table 2). The reactions were of first order with respect to the concentration of the substrate (Fig. 1). The reactions were catalyzed by acid, and the catalyzed reaction showed a first-order dependence on acidity (Fig. 2). The linear increase in the oxidation rate with acid concentration indicated the involvement of a protonated Cr(VI) species in the rate-determining step of the reaction. Earlier reports have established the involvement of protonated Cr(VI) species in chromic acid oxidation reactions.[4,9] These data suggest the following rate law:

$$v = k[\text{Substrate}][\text{QDC}][\text{H}^+] \qquad (2)$$

Table 1. Stoichiometry of the Oxidation of Substrates by QDC (Concentration of Substrate=0.005 M)

Parameter	[H_2SO_4] (M) 0.10	0.25	0.50
10^2[QDC]/M	2.50	2.60	2.70
Δ[QDC]/Δ[Borneol]	0.31	0.33	0.34
Δ[QDC]/Δ[Isoborneol]	0.37	0.35	0.33

Table 2. Dependence of Rate Constant on Oxidant Concentration at 323 K ([Substrate] 0.01 M; [H_2SO_4] 0.02 M)

[QDC] $10^3\times$ M	Borneol $10^4k_1/\text{s}^{-1}$	Isoborneol $10^4k_1/\text{s}^{-1}$
0.10	1.15	2.84
0.25	1.09	2.75
0.50	1.02	2.88
0.75	1.13	2.72
1.0	1.07	2.81

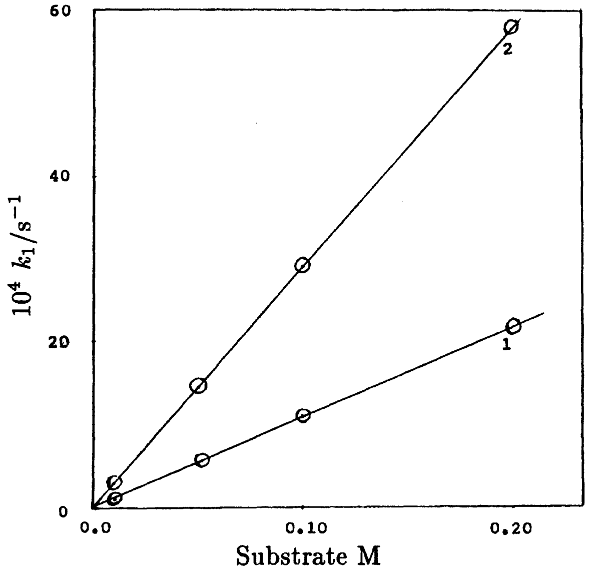


Fig. 1. Plot of k_1 against substrate concentration (1: Borneol; 2: Isoborneol).

Effects of Solvent. The acid-catalyzed oxidation of the substrates was studied in solutions containing varying proportions of DMF and water. Plots of $\log k_1$ against the inverse of the dielectric constant were linear (r=0.992) with positive slopes (Fig. 3). This suggested an ion–dipole interaction,[10] in agreement with the observation that in the presence of an acid, the rate-determining step involved a protonated Cr(VI) species.

Effects of Temperature. The catalyzed oxidation of the substrates was studied at different temperatures and it was observed that the rate increased with an increase in temperature (Table 3). From the linear plots of $\log k_1$ against the reciprocal of temperature, the activation energy was calculated. The entropy of activation ($\Delta S^{\neq}$), and the free energy of activation ($\Delta G^{\neq}$), were calculated using the Eyring equation and the Gibbs equation, respectively.[11]

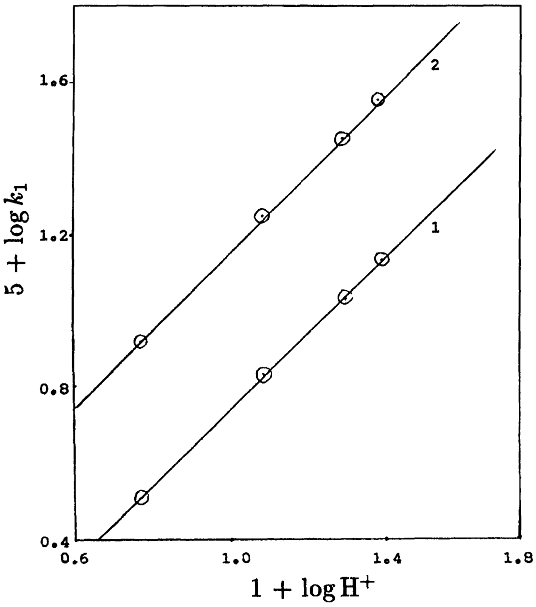


Fig. 2. Plot of $\log k_1$ against $\log \text{H}^+$ (1: Borneol; 2: Isoborneol).

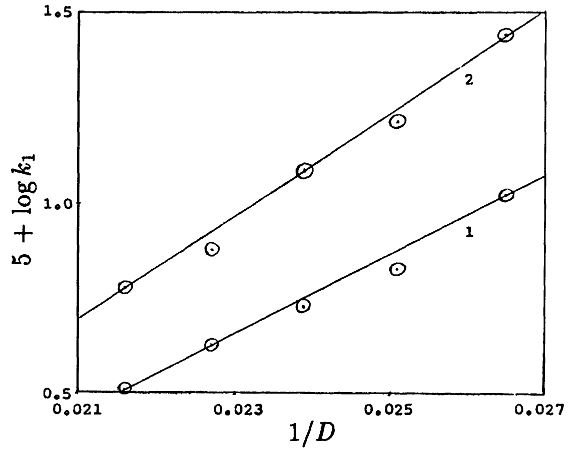


Fig. 3. Plot of $\log k_1$ against reciprocal of dielectric constant (1: Borneol; 2: Isoborneol).

Table 3. Dependence of Rate Constant ($10^4k_1/\mathrm{s}^{-1}$) on Temperature[a)] and Activation Parameters[b)]

Temp (±0.1 K)	Borneol	Isoborneol
313	0.46	1.31
318	0.69	1.99
323	1.07	2.81
328	1.60	3.32
333	2.36	3.84
$\Delta H^{\neq}$ (kJ mol^{-1})	68	25
$\Delta S^{\neq}$ (J K^{-1}mol^{-1})	−111	−238
$\Delta G^{\neq}$ (kJ mol^{-1})	104	102

a) [Substrate] 0.01 M; [QDC] 0.001 M; [H_2SO_4] 0.02 M.
b) Error limits: $\Delta H^{\neq}$ ±2 kJ mol^{-1}; $\Delta S^{\neq}$ ±3 J K^{-1} mol^{-1}; $\Delta G^{\neq}$ ±2 kJ mol^{-1}.

Kinetic Isotope Effect. The rate of oxidation of isoborneol-2-*d* has been measured under experimental conditions identical with that used for the oxidation of isoborneol. The rates of oxidation of isoborneol and isoborneol-2-*d* at 323 K were 10^4k_1 (s^{-1})=28.1 and 4.46, respectively. The kinetic isotope effect, k_H/k_D, was 6.3 at 323 K.

Induced Polymerization. It was observed that there was no induced polymerization of acrylonitrile or the reduction of mercury(II) chloride.[12)] This indicated that a one-electron oxidation was quite unlikely. Control experiments done in the absence of the substrates did not show any appreciable change in the concentration of QDC.

Relative Rates of Oxidation. The rates of oxidation of borneol and isoborneol have been compared with those of α- and β-norborneols and cyclohexanol, and the results are given in Table 4.

Mechanism. An examination of the kinetic data obtained in this investigation showed that the rate of oxidation of the alcohols was dependent on the first powers of the concentrations of each — substrate, oxidant, and acid. The acid catalysis of the reaction must be related to the structure of the oxidant (QDC) which was converted into a protonated species, even at the low concentrations of mineral acid used.

The first serious attempt to understand the mechanism of chromic acid oxidation of alcohols was made by Westheimer, who established that the oxidation proceeded via the intermediate formation of a chromate,[13—15)] which then underwent decomposition to the corresponding carbonyl compound.[16)] The question of the intermediacy of the chromic ester was settled by an investigation of the oxidation of sterically hindered alcohols by chromic acid.[17)]

Table 4. Rate Data for the Oxidation of Some Secondary Alcohols[a)] at 323 K

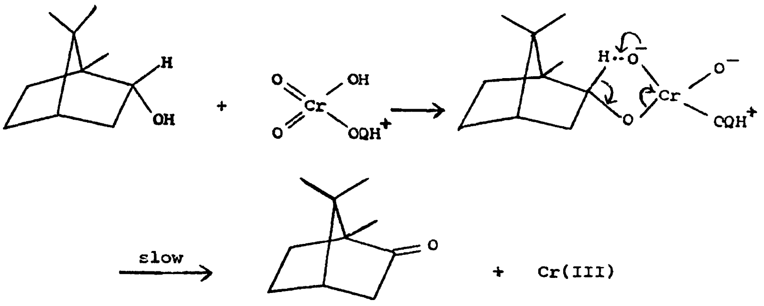
Alcohol	$10^4k_1/\mathrm{s}^{-1}$	k_{relative}
Borneol (*endo*)	1.07	3.7
Isoborneol (*exo*)	2.81	9.7
α-Norborneol (*endo*)	0.61	2.1
β-Norborneol (*exo*)	0.41	1.4
Cyclohexanol	0.29	1.0

a) [Alcohol] 0.01 M; [QDC] 0.001 M; [H_2SO_4] 0.02 M.

Our data on the oxidation of secondary alcohols by QDC seem to suggest that if the strain-induced instability of the chromate were to decide the relative rates of oxidation, then there would be a rate-accelerating effect resulting from the relief of steric strain in the chromate. Our results would help to understand the extent of the influence of strain relief in deciding the rate of chromate decomposition. For this purpose, the effects of the strain relief factor on the relative rates of oxidation of alcohols derived from bicyclo [2.2.1] heptane by QDC have been examined. The results in Table 4 show that borneol, with an axial hydroxy group, is more stable than the isomeric equatorial isoborneol. In isoborneol, the chromate substituent group is in the *exo* configuration and is more strained due to the bridge *gem*-dimethyl groups. The hydroxy group suffers nonbonded repulsion from the *gem*-dimethyl group at C-7, this effect being diminished when the transition state is formed. There is a rate-enhancing relief of steric strain in the transition state, and isoborneol is oxidized more rapidly than borneol by a factor of 2.63. This indicates a difference in the ease of ester decomposition. Such a conclusion is supported by the results obtained for the oxidation of norborneols. When their reactivities towards oxidation by QDC are considered, α-norborneol (*endo*) is more reactive than β-norborneol (*exo*) by a factor of 1.5 (Table 4). α-Norborneol (*endo*) forms the chromate with greater difficulty and is more unstable than β-norborneol (*exo*) due to the strain induced by nonbonded interaction with the axial transannular hydrogen.[18)]

The difference in the rates of oxidation between isoborneol and α-norborneol (*endo*) gives a factor of 4.6, which suggests that the deciding factor in the decomposition of the chromate is related to nonbonded repulsions, which contribute to the instability of the chromic ester. The same argument would be valid for the series, α-norborneol > cyclohexanol (Table 4). The factor of 2.1 that distinguishes the oxidation rates of α-norborneol and cyclohexanol is a pointer to the strain and steric hindrance developed by substitution in the axial position of the boat structure of α-norborneol.

A stoichiometric conversion of the secondary alcohols (borneol and isoborneol) to the corresponding ketone (camphor) was observed. This oxidation process involved the change $\rangle\mathrm{CHOH}\longrightarrow\rangle\mathrm{C{=}O}$, and the reaction resembled the oxidation of other secondary alcohols.[19,20)] If the reaction intermediate were to be visualized as having a cyclic structure, then this would explain all the features of the oxidation reaction. The negative entropy of activation would be consistent with the formation of a cyclic transition state in a bimolecu-



Scheme 1.

lar reaction. The similarity of $\Delta G^{\neq}$ values arises from changes in $\Delta H^{\neq}$ and $\Delta S^{\neq}$ values, and stresses the probability that these oxidation reactions involve similar rate-determining steps. If the chromium were to be coordinated through the alcohol O–H group, then the formation of the chromate would be facilitated. This would increase the ease of oxidation of the alcohols (borneol and isoborneol), and their conversion to the ketone (camphor) could then be rationalized.

If there had been no kinetic isotope effect (that is, if k_H/k_D had been equal to unity), then the formation of the chromic ester would have been the rate-determining step and not the cleavage of the alcohol carbon–hydrogen bond. Since the kinetic isotope effect was observed to be $k_H/k_D=6.3$, this would establish the cleavage of the carbon–hydrogen bond as the rate-determining step of the reaction.[21] Generally, for a cyclic mechanism with a rate-determining carbon–hydrogen bond cleavage, the k_H/k_D values were fairly high, as seen from the oxidation of alcohols (both primary and secondary) by various oxidants.[21—24]

In this study, a cyclic transition state is thus favored. The proton is removed in the cyclic transition state, the center of which resides on an electron-dense oxygen on the chromate.[25] In such situations, steric hindrance afforded by groups near the reaction center could result in an acceleration of the rate of the reaction, by restricting the rotation about the C–O bond. There would be an increased probability of attaining the five-membered cyclic transition state, in which there would be coplanarity of all the atoms involved. Such a transition state would mean the transfer of electrons towards the chromium occurring by the formation of the carbon–hydrogen–oxygen bonds as well as the carbon–oxygen–chromium bonds. Partly occupied orbitals are thus used to bind the transferred hydrogen to both carbon and oxygen in the transition state (Scheme 1). A cyclic transition state has been suggested as the most plausible intermediate for the oxidation of secondary alcohols by Cr(VI) reagents.[26] Moreover, this electrocyclic mechanism for the oxidation of bicyclic alcohols by QDC involves six electrons, and being a Hückel-type system $(4n+2)$, this would be an allowed process.[27]

It had been earlier suggested that the oxidation of alcohols could also proceed by way of an intermediate formed by removing a hydride ion from the hydroxyl group of the alcohol.[28,29] The rate-determining step would have been the formation of this intermediate. Clearly, no such intermediate could occur in the oxidation of bicyclic alcohols under our experimental conditions. The kinetic isotope effect, $k_H/k_D=6.3$, observed in this investigation, ruled out this possibility.

Financial support from the University Grants Commission, New Delhi, under the Special Assistance Program, is gratefully acknowledged.

References

1) H. G. Kuivila and W. J. Becker, *J. Am. Chem. Soc.*, **74**, 5329 (1952).
2) S. Winstein and N. J. Holness, *J. Am. Chem. Soc.*, **77**, 5562 (1955).
3) H. Kwart and P. S. Francis, *J. Am. Chem. Soc.*, **77**, 4907 (1955).
4) K. B. Wiberg, "Oxidation in Organic Chemistry," Academic Press, New York (1965), Part A, p. 69.
5) A. K. Awasthy, J. Rocek, and R. M. Moriarty, *J. Am. Chem. Soc.*, **89**, 5400 (1967).
6) I. Nongkynrih and M. K. Mahanti, *J. Org. Chem.*, **58**, 4925 (1993).
7) K. B. Wiberg and H. Schafer, *J. Am. Chem. Soc.*, **91**, 927 (1969).
8) K. Balasubramanian and V. Prathiba, *Indian J. Chem., Sect. B*, **25B**, 326 (1986).
9) K. K. Banerji, *J. Chem. Res., Miniprint*, **1978**, 2561; *Indian J. Chem., Sect. A*, **17A**, 300 (1979).
10) E. S. Amis, "Solvent Effects on Reaction Rates and Mechanisms," Academic Press, New York (1967), p. 42.
11) A. A. Frost and R. G. Pearson, "Kinetics and Mechanism," Wiley Eastern, New Delhi (1970), p. 99.
12) J. S. Littler and W. A. Waters, *J. Chem. Soc.*, **1959**, 1299.
13) F. H. Westheimer and A. Novick, *J. Chem. Phys.*, **11**, 506 (1943).
14) F. H. Westheimer, *Chem. Rev.*, **45**, 419 (1949).
15) W. Watanabe and F. H. Westheimer, *J. Chem. Phys.*, **17**, 61 (1949).
16) F. Holloway, M. Cohen, and F. H. Westheimer, *J. Am. Chem. Soc.*, **73**, 65 (1951).
17) J. Rocek, F. H. Westheimer, A. Eschenmoser, L. Moldovanyi, and J. Schreiber, *Helv. Chim. Acta*, **45**, 2554 (1962).

18) D. H. R. Barton, *J. Chem. Soc.*, **1953**, 1027.

19) M. Cohen and F. H. Westheimer, *J. Am. Chem. Soc.*, **74**, 4387 (1952).

20) J. Rocek and J. Krupicka, *Collect. Czech. Chem. Commun.*, **23**, 2068 (1958).

21) F. H. Westheimer and N. Nicolaides, *J. Am. Chem. Soc.*, **71**, 25 (1949).

22) J. S. Littler and W. A. Waters, *J. Chem. Soc.*, **1959**, 4046.

23) K. K. Banerji, *Bull. Chem. Soc. Jpn.*, **51**, 2732 (1978).

24) R. A. Pascal, Jr., M. W. Baum, C. K. Wagner, L. R. Rodgers, and D.-S. Huang, *J. Am. Chem. Soc.*, **108**, 6477 (1986).

25) R. T. Arnold and M. J. Danzig, *J. Am. Chem. Soc.*, **79**, 892 (1957).

26) S. Agarwal, H. P. Tiwari, and J. P. Sharma, *Tetrahedron*, **46**, 1963 (1990).

27) J. S. Littler, *Tetrahedron*, **27**, 81 (1971).

28) J. Rocek and J. Krupicka, *Chem. Ind.*, **1957**, 1668.

29) J. Rocek, *Collect. Czech. Chem. Commun.*, **25**, 1052 (1960).