Permanganate Oxidations. III. Kinetics and Mechanisms of the Oxidation of Furfurals in Alkaline Media[1-3]

Fillmore Freeman,[4] Jay B. Brant, Norman B. Hester, Albert A. Kamego, Marc L. Kasner, Thomas G. McLaughlin, and Edward W. Paull

Department of Chemistry, California State College, Long Beach, California 90801

A spectrophotometric stopped-flow kinetic study of the permanganate ion oxidation of furfural (I) and six 5-substituted furfurals at pH 11.5–13.3 reveals that the reaction follows two reaction paths. The minor pathway (Scheme I) is independent of hydroxyl ion concentration, and the major mechanism (Scheme II) is dependent on the first power of hydroxide ion concentration. Both reaction pathways are first order with respect to the concentration of I and permanganate ion. A correlation of the second-order rate constants with Hammett σ *meta*-substituent constants has been observed for the substituents 5-Me, 5-Et, 5-n-Bu, H, 5-Cl, and 5-Br at 25° with $\rho = +1.30$ (Scheme II). At pH 13.3 (Scheme II), $\Delta H \neq$ is 10.2 kcal/mol, $\Delta S \neq$ is -22.8 eu, and k_H/k_D is >1.8. Oxygen-18 experiments show that the solvent is the major source of oxygen introduced into I *via* Scheme II. The kinetic data are consistent with the formation of the hydrate anion of I followed by a hydride anion transfer to permanganate ion in the rate-determining step for the mechanism of Scheme II. It is postulated that the mechanism of Scheme I involves a direct attack of permanganate ion on I to give the permanganate ester, which decomposes in a subsequent slow step.

Although a few kinetic studies of the permanganate ion oxidation of aliphatic and aromatic aldehydes have been published,[5-7] no reports have appeared concerning the oxidation of heterocyclic aldehydes. Since there is the possibility of simultaneous attack at the furan ring and at the carbonyl group, several conflicting reports have appeared concerning the permanganate ion oxidation of furfural (I). For example, Obata,[8] using an excess of oxidant, obtained maleic acid from the alkaline permanganate oxidation of I. In contrast, Wagner and Simons[9] reported a greater than 80% yield of 2-furoic acid (II) from I, and Gilman and Wright[10] oxidized 5-bromofurfural (III) to the corresponding

$$I, X = H$$
$$III, X = Br$$

$$II, X = H$$

$$(1)$$
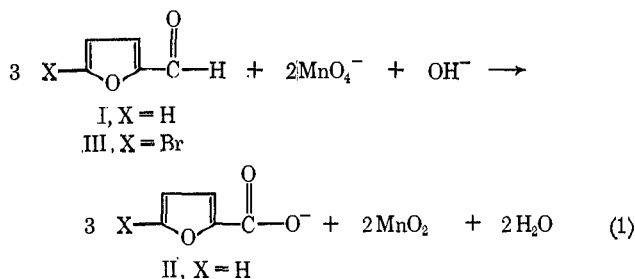
acid in basic media. The complexity and pH dependence of the permanganate ion oxidation of I are further demonstrated by the observations that II is not the major oxidation product in neutral and acid solution.[11,12] Presumably, the furan ring is more susceptible to attack than the carbonyl carbon in nonalkaline media.

Owing to the absence of previous kinetic studies, we have investigated the alkaline permanganate ion oxidation of I spectrophotometrically in a stopped-flow reactor.[1,13] A study of the effect of substituents on the rates of reaction, of kinetic isotope effects, of activation parameters, and of oxygen-labeling experiments permits a fairly clear mechanistic picture to be presented. Also, one of the few examples of an application of the Hammett equation to the furan nucleus is discussed.[14-17]

Experimental Section

All melting points are uncorrected and were determined on a Thomas–Hoover apparatus. Elemental analyses were performed by Spang Microanalytical Laboratory, Ann Arbor, Mich. Mass spectral analyses were performed by Professor P. C. Ford.[18]

Reagents.—Distilled water was purified by passing through an ion-exchange cartridge (Type R-2, Illinois Water Treatment Company, Rockford, Ill.). Reagent grade sodium nitrate (Baker) was used to adjust solutions to desired ionic strengths. Potassium permanganate stock solutions were prepared from Acculute standard volumetric solutions. The stock solution was stored under nitrogen and the absorbancy index of permanganate ion was checked before each set of kinetic runs. Sodium hydroxide solutions were prepared from standard concentrated (Acculute). All solutions were prepared immediately before the kinetic runs, and the pH of the solutions were taken as those measured potentiometrically.

Potassium permanganate-^{18}O was prepared by the isotopic exchange reaction between normal permangatate ion (Mallinckrodt) and water-[18]O (Bio-Rad).[5,19] Isotopic analyses were performed by thermal decomposition of the sample followed by mass spectrometric analysis of the oxygen formed.[5,20] The m/e 34:32 ratio was used to calculate per cent [18]O.

Furfural (Aldrich) was purified by distillation, bp 91–93° (11 mm).

(1) Part II: F. Freeman, A. Yeramyan, and F. Young, *J. Org. Chem.*, **34**, 2438 (1969).

(2) Abstracted in part from the M.S. thesis of E. W. Paull, California State College, Long Beach, 1969.

(3) Presented in part before the Symposium on the Oxidation of Organic Compounds at the American Chemical Society Pacific Conference on Chemistry and Spectroscopy, San Francisco, Calif., Nov 8, 1968.

(4) To whom correspondence should be addressed.

(5) K. B. Wiberg and R. Stewart, *J. Amer. Chem. Soc.*, **77**, 1786 (1955).

(6) R. Stewart and M. M. Mocek, *Can. J. Chem.*, **41**, 1160 (1963).

(7) J. Rocek in "The Chemistry of the Carbonyl Group," S. Patai, Ed., Interscience Publishers, Inc., New York, N. Y., 1966, p 478.

(8) Y. Obata, *J. Agr. Chem. Soc. Jap.*, **16**, 187 (1940); *Chem. Abstr.*, **34**, 5840 (1940).

(9) E. C. Wagner and J. K. Simons, *J. Chem. Educ.*, **13**, 265 (1936).

(10) H. Gilman and G. F. Wright, *J. Amer. Chem. Soc.*, **52**, 1170 (1930).

(11) F. Freeman, unpublished data, 1968.

(12) A. P. Dunlop and F. N. Peters, "The Furans," Reinhold Publishing Corp., New York, N. Y., 1953, pp 385, 494, 496.

(13) K. B. Wiberg and R. D. Geer, *ibid.*, **87**, 5202 (1965); **88**, 5827 (1966).

(14) The first application of the Hammett law to the furan ring was with 2-furoic acids.[15] Noyce and Kaiser[16] obtained a ρ value of -7.5 for the solvolysis of 5-substituted furylmethylcarbinyl p-nitrobenzoates. We express our thanks to Professor Noyce for providing us with a copy of their manuscript prior to publication.

(15) E. Imoto and R. Motoyama, *Bull. Naniwa Univ.*, *Ser. A*, **2**, 127 (1954).

(16) D. S. Noyce and G. V. Kaiser, *J. Org. Chem.*, **34**, 1008 (1969).

(17) (a) F. Freeman, *J. Chem. Educ.*, in press. (b) H. Van Bekkum, P. E. Verkade, and B. M. Wepster, *Rec. Trav. Chim. Pays-Bas*, **78**, 815 (1959). (c) R. W. Taft, *J. Phys. Chem.*, **64**, 1805 (1960).

(18) West Coast Technical Service, San Gabriel, Calif.

(19) G. L. Zimmerman, Ph.D. Thesis, University of Chicago, 1949.

(20) T. C. Hoering and J. W. Kennedy, *J. Amer. Chem. Soc.*, **78**, 4829 (1956).

5-Nitrofurfural was prepared according to the method of Gilman and Wright,[21] using 5-nitrofurfural diacetate (Eastman). Under an atmosphere of nitrogen, 24.3 g (0.1 mol) of 5-nitrofurfural diacetate was refluxed with a solution of 54 g (0.56 mol) of concentrated sulfuric acid in 108 g (6.0 mol) of water for 15 min with constant stirring. The ether extract of the cooled hydrolyzed product was washed with water, dried over Na$_2$SO$_4$, and filtered. The ether was removed and the residue was distilled at 132–134° (11 mm) to give 10 g (71%) of product, mp 34–36° (lit.[21] mp 35–36°).

5-Methylfurfural was prepared from 2-methylfuran (Aldrich) according to the method of Traynelis, et al.,[22] except that a ratio of 2 mol of DMF and 2 mol of POCl$_3$ to 1 mol of 2-methylfuran was used. Distillation at 78–80° (16 mm) [lit.[22] bp 72–73° (13 mm)] gave 88 g (76.1%) of product.

5-Ethylfurfural was prepared from 2-ethylfuran, which was prepared by reduction[23,24] of 2-acetylfuran (Aldrich). 2-Ethylfuran was formylated as described above. Distillation gave 11.9 g (64%) of product, bp 88–90° (11 mm) [lit.[22] bp 88–90° (11 mm)].

5-Isobutylfurfural was prepared from 2-isobutylfuran, which was prepared by reduction of isobutyl furyl ketone. Isobutyl furyl ketone was prepared by the Friedel–Crafts reaction of furan and isobutyryl chloride.[24]

In a 1-l., round-bottomed flask, 38.8 g (0.28 mol) of isobutyl furyl ketone, 28 ml (0.5 mol) of 85% hydrazine hydrate, 28 g (0.7 mol) of sodium hydroxide, and 225 g (3.64 mol) of ethylene glycol were heated for 2.5 hr until the vigorous evolution of gas diminished. A two-phase distillate was obtained after distillation. The aqueous layer was separated and extracted twice with 50-ml portions of ether. The combined organic solution was dried (Na$_2$SO$_4$), filtered, and distilled to give 14.3 g (41.2%) of 2-isobutylfuran, bp 126.5–128° (lit.[24] bp 123–127°). Formylation of the product as described above gave 13.6 g (72%) of 5-isobutylfurfural, bp 168–170° (40 mm). The 2,4-dinitrophenylhydrazone, mp 150–152°, was prepared and analyzed.

Anal. Calcd for C$_{15}$H$_{16}$N$_4$O$_5$: C, 54.22; H, 4.85; N, 16.86. Found: C, 54.05; H, 4.90; N, 16.76.

5-*n*-Butylfurfural was prepared from 2-*n*-butylfuran. *n*-Butyl furyl ketone, which was prepared by the Friedel–Crafts reaction of furan and *n*-butyryl chloride,[24] was reduced as described above[23,24] to give 2-*n*-butylfuran, bp 140–142° (lit.[24] bp 137–138°). Formylation gave 50.1 g (91%) of 5-*n*-butylfurfural, bp 117–119° (11 mm). The 2,4-dinitrophenylhydrazone, mp 160–162°, was analyzed.

Anal. Calcd for C$_{15}$H$_{16}$N$_4$O$_5$: C, 54.22; H, 4.85; N, 16.86. Found: C, 54.00; H, 4.86; N, 16.75.

5-Chlorofurfural was prepared, mp 31–33° (lit.[25] mp 31.5–33°).

5-Bromofurfural was prepared, mp 83–85° (lit.[26,27] mp 83–85°).

Furfural-*d$_1$* was prepared in 85% isotopic purity (pmr) by the Rosenmund reduction[28] of 2-furoyl chloride.[29] Attempts to prepare furfural-*d$_1$* by the lithium deuteride reduction of furil[30] or the manganese dioxide oxidation of α,α-*d$_2$*-furfuryl alcohol[31] were unsuccessful.

Since furfurals are easily oxidized, they were stored as the bisulfite addition compound. The only exception was the water-soluble bisulfite addition compound of I.

Oxidation with Potassium Permanganate-^{18}O.—A modification of the method of Wiberg and Stewart[5] was used to prepare 5-bromofuroic acid for mass spectral analysis. In a 2-l. flask fitted with stirrer, thermometer, and nitrogen inlet flow, which contained 1.28 g (0.0072 mol) of 5-bromofurfural in 750 ml of de-

ionized water equilibrated in a water bath at 25.0°, was rapidly added 250 ml of deionized water containing 0.76 g (0.0048 mol) of labeled potassium permanganate (0.833% ^{18}O) and 150 ml of 2 *N* sodium hydroxide (added just prior to addition). The resulting solution had a pH of 13.3. After 5 min the solution was quenched with 50 ml of 0.5 *M* sodium bisulfite, concentrated sulfuric acid was added to dissolve precipitated manganese dioxide, and the solution was extracted three times with 75-ml portions of benzene. The extract was washed with water and extracted with 40 ml of 0.5 *M* potassium carbonate solution. Acidification of the basic solution with concentrated sulfuric acid gave, after cooling, filtering, and drying, 0.24 g of 5-bromofuroic acid, mp 183–185° (lit.[32] mp 185–186°). The ^{18}O content of the acid was analyzed and the per cent transfer of oxygen from potassium permanganate was calculated according to the formula of Wiberg and Stewart:[5,33] % O from KMnO$_4$ = 100 [2(*y* − 0.204)/*x* − 0.204], where *y* = 100 [^{18}O/(^{16}O + ^{18}O)] and *x* = %.^{18}O in KMnO$_4$ = 0.833%.

Kinetic Method.—All experiments were performed under pseudo-first-order conditions using a large excess of furfural.

Because of the large rate constants, the rates of reaction were determined by following spectrophotometrically the disappearance of permanganate ion (522 mμ) in a stopped-flow system.[1,13] The rate constants, which were obtained from plots of −ln[log (T_∞/T)] *vs.* time, where T_∞ is the per cent transmission at a point just before colloidal manganese dioxide begins to form, were calculated on an IBM 1620 computer.[34] The rates were followed until the reactions were 75–90% complete, and the rate constants were calculated three times using the data to the first half-life, to the second half-life, and then to the third half-life. In this method, any deviation from linearity was readily observed.

Results and Discussion

The rate of oxidation of I was determined using an eightfold range of furfural concentration, a threefold range of permanganate ion concentration, and a sevenfold range of hydroxide ion concentration (Tables I and II). A plot of k_ψ (pseudo-first-order rate constant) *vs.* concentration of I gives a straight line that goes through the origin, indicating the rate of oxidation to have a first-order dependence on furfural concentration. At constant furfural and hydroxide ion concentrations, the pseudo-first-order rate constant is not altered appreciably with increasing permanganate ion concentrations, which indicates a first-order dependence on permanganate ion. With a large excess of I and hydroxide ion, good first-order plots are obtained, which further demonstrates that the reaction has a first-order dependence on permanganate ion. A plot of the second-

(21) H. Gilman and F. Wright, *J. Amer. Chem. Soc.*, **52**, 2550 (1925).

(22) J. T. Traynelis, J. Miskel, Jr., and J. R. Sawa, *J. Org. Chem.*, **22**, 2170 (1957).

(23) A. K. Armitage and H. R. Ing, *Brit. J. Pharmacol.*, **9**, 376 (1954); *Chem. Abstr.*, **49**, 10918 (1955).

(24) H. Gilman and N. O. Calloway, *J. Amer. Chem. Soc.*, **55**, 4197 (1933).

(25) W. J. Chute and F. Wright, *J. Org. Chem.*, **10**, 541 (1945).

(26) Z. N. Nazarova, *Zh. Obshch. Khim.*, **24**, 575 (1954); *Chem. Abstr.*, **49**, 6214 (1955).

(27) Z. N. Nazarova, *Dokl. Akad. Nauk Uz. SSR*, No. 4, 40 (1953); *Ref. Zh., Khim.*, No. 39483 (1954); *Chem. Abstr.*, **49**, 10261 (1955).

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(29) G. S. Saharia, *J. Sci. Ind. Res.*, **15B**, 69 (1954); *Chem. Abstr.*, **50**, 14705 (1956).

(30) K. B. Wiberg, *J. Amer. Chem. Soc.*, **76**, 5371 (1954).

(31) E. F. Pratt and J. F. Van de Castle, *J. Org. Chem.*, **26**, 2973 (1961).

TABLE I

RATE DEPENDENCE ON FURFURAL AND PERMANGANATE CONCENTRATIONS[a]

[Furfural], $10^2 M$	[OH$^-$], M	[MnO$_4^-$], $10^4 M$	k_ψ,[b] sec^{-1}	k_2,[c] M^{-1} sec^{-1}
0.25[d]	0.20	4.0	0.02	8.0
0.50[d]	0.20	4.0	0.04	8.0
2.01[d]	0.20	4.0	0.159	7.9
4.02[d]	0.20	4.0	0.34	8.4
4.02	0.20	4.0	0.54	13.4
4.02	0.20	6.0	0.60	14.9
4.02	0.20	8.0	0.55	13.7
4.02	0.20	12.0	0.65	16.1

[a] λ = 522 mμ, μ = 1.0 *M*, *T* = 25.0°. [b] Pseudo-first-order rate constant. [c] Second-order rate constant = k_ψ/[furfural]. [d] *T* = 5.0°.

(32) P. S. Bailey and J. V. Waggoner, *ibid.*, **15**, 159 (1950).

(33) K. B. Wiberg and R. Stewart, *J. Amer. Chem. Soc.*, **78**, 1214 (1956).

(34) Professor K. B. Wiberg's program for obtaining rate constants for first-order reactions was used: K. B. Wiberg, "Computer Programming for Chemists," W. A. Benjamin, Inc., New York, 1965, p 168 ff.

TABLE II

RATE DATA FOR THE NEUTRAL AND BASE-CATALYZED REACTIONS OF
FURFURAL AND 5-METHYL- AND 5-NITROFURFURAL[a]

Substrate	pH	$[OH^-]$, M	k_ψ,[b] sec^{-1}	k_2,[c] $M^{-1} sec^{-1}$	k_0,[d] $M^{-1} sec^{-1}$	$k_3 = (k_2 - k_0)/[OH^-]$,[e] $1.^2 mol^{-2} sec^{-1}$
I[f]	12.5	0.03	0.18	4.50	2.70	60.0
I[f]	12.8	0.063	0.26	6.50	2.70	60.3
I[f]	13.25	0.18	0.54	13.50	2.70	60.0
I[f]	13.3	0.20	0.59	14.75	2.70	60.3
IV[g]	12.3	0.02	0.14	3.50	2.60	45.0
IV[g]	12.5	0.03	0.15	3.75	2.60	38.3
IV[g]	12.8	0.063	0.20	5.00	2.60	38.1
IV[g]	13.25	0.18	0.39	9.75	2.60	39.7
V[h]	11.5	0.003	0.61	127.6[i]	47.8	26.6×10^3
V[h]	11.8	0.006	1.03	215.5[i]	47.8	27.9×10^3
V[h]	12.3	0.02	2.74	573.2[i]	47.8	26.3×10^3

[a] $\lambda = 522$ mμ, $T = 25.0°$, $\mu = 1.0 M$. [b] Pseudo-first-order rate constant. [c] Second-order rate constant $= k_\psi/[\text{Furfural}]$. [d] Intercept from plot of k_2 vs. $[OH^-]$. [e] Third-order rate constant which agrees well with the slope from a plot of $(k_2 - k_0)$ vs. $[OH^-]$. [f] $[I] = 4.02 \times 10^{-2} M$. [g] $[IV] = 4.00 \times 10^{-2} M$. [h] $[V] = 4.78 \times 10^{-3} M$. [i] $T = 14.5°$.

order rate constant $(k_2 = k_\psi/[\text{furfural}])$ vs. hydroxide ion concentration gives a straight line that does not go through the origin. Similar plots were obtained with 5-methylfurfural (IV) and 5-nitrofurfural (V). This implies that there are both zero-order and first-order terms describing the effect of hydroxyl ion on the rate of reaction. These data suggest the following rate law

$-d[MnO_4^-]/dt =$

$k_0[\text{furfural}] [MnO_4^-] + k_3[\text{furfural}] [MnO_4^-][OH^-]$ (2)

where k_0 (the intercept of a plot of k_2 vs. $[OH^-]$) is the rate constant for oxidation of neutral I, IV, or V, and k_3 is the constant for the base-catalyzed reaction.[35] Table II shows the constancy of k_3 for I, IV, and V, respectively.

A possible mechanism for the neutral reaction, which involves a direct attack of permanganate ion at the carbonyl carbon of I to give the permanganate ester V, is shown in Scheme I. The slow step is probably

SCHEME I

$3MnO_3^- + H_2O \xrightarrow{k_4}_{\text{fast}} 2MnO_2 + MnO_4^- + 2OH^-$
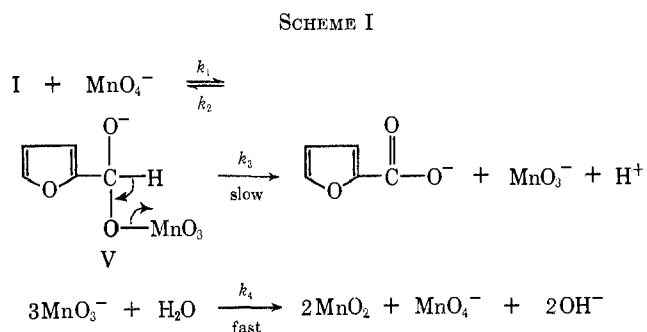
the decomposition of V via proton transfer and cleavage of the manganese–oxygen bond. The rate equation derived from this mechanism is eq 3, where K_{eq} denotes the equilibrium constant for the formation of V.[7]
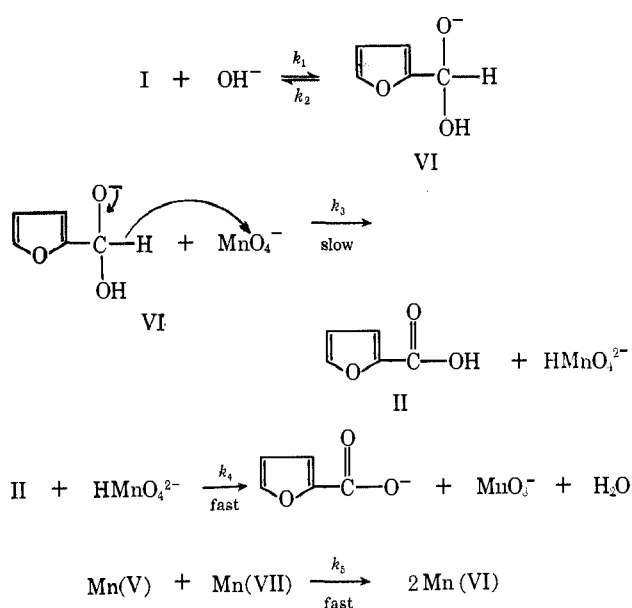
$\nu = k_\psi[IV] = k_0 K_{eq}[I][MnO_4^-]$ (3)

The mechanism shown in Scheme II, for the base-catalyzed reaction, presumably involves a hydride transfer in the rate-determining step. The rate equation derived from this mechanism is eq 4, where K_{eq} is the equilibrium constant for the formation of VI.

$\nu = k_\psi[VI] [MnO_4^-] = k_{OH} \cdot K_{eq}[I][MnO_4^-][OH^-]$ (4)

SCHEME II

$Mn(V) + Mn(VII) \xrightarrow{k_5}_{\text{fast}} 2 Mn(VI)$

The mechanism correctly predicts that the rate will be a linear function of hydroxide ion concentration and that the reaction will be subject to specific hydroxide ion catalysis.

The observed rate of disappearance of permanganate ion is the sum of the rates of the neutral (Scheme I) and base-catalyzed (Scheme II) reactions. This means that a plot of $(k_2 - k_0)$ vs. hydroxide ion concentration should give a straight line passing through the origin if the reaction is first order in hydroxyl ion. Using the data from Table II shows that this is indeed true for I, IV, and V.

Scheme I suggests that the permanganate ion will be the source of oxygen introduced into I, while Scheme II predicts that the oxygen will come from the solvent. In order to determine the source of oxygen, 5-bromofurfural (III)[36] was oxidized with enriched potassium permanganate-^{18}O and then analyzed via mass spectrometry (Table III). The low value at pH 13.3 and the larger value at pH 12.3 support Scheme II. The low incorporation of ^{18}O into I, at pH 13.3, could arise

(35) Our results in neutral and acid media indicate that k_0 might also contain the rate constant for permanganate attack on the furan nucleus.[11]

(36) Compound III was selected because 5-bromo-2-furoic acid is less water soluble than 2-furoic acid (II).

TABLE III
OXIDATION OF 5-BROMOFURFURAL WITH
POTASSIUM PERMANGANATE-^{18}O AT 25°

pH	% O from MnO$_4^-$
13.3	7.6
12.3	21.3

from the slow exchange between permanganate ion and solvent.[37-41]

The enthalpy of activation (10.2 kcal/mol) and the entropy of activation (−22.8 eu) at pH 13.3 for the base-catalyzed oxidation of I are of the same order of magnitude as the values for the reaction of other anions with permanganate.[1,42,43]

Furfural-d_1 was prepared in 85% isotopic purity via the Rosenmund reduction of furoyl chloride (eq 5).[28,29,44]

$$\text{(furoyl chloride)}\!-\!\overset{\text{O}}{\underset{\|}{\text{C}}}\!-\!\text{Cl} \;+\; \text{D}_2 \;\xrightarrow{\text{Pd--S}}\; \text{(furfural-}d_1\text{)}\!-\!\overset{\text{O}}{\underset{\|}{\text{C}}}\!-\!\text{D} \qquad (5)$$
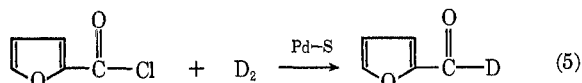
An uncorrected k_H/k_D value of 1.8 was observed at pH 13.3. Although no quantitative significance can be ascribed to this value, the results do support the breaking of a C–H(D) bond in the rate-determining step (Scheme II).

Table IV shows a slight increase in reaction rate owing to increased salt concentration, which is expected for a reaction between two negative ions.[46]

TABLE IV
KINETIC DEPENDENCE ON IONIC STRENGTH AT 25.0°[a]

μ	k_ψ,[b] sec^{-1}
0.1	0.37
0.2	0.42
0.3	0.43
0.5	0.51

[a] [Furfural] = 4.02 × 10^{-2} M, [MnO$_4{}^{2-}$] = 4.00 × 10^{-4} M, pH = 13.3, λ = 522 mμ. [b] Pseudo-first-order rate constant.

Since the furan ring is a planar pentagon with sp^2-hybridized carbon atoms and possesses considerable aromatic character arising from the delocalization of the two paired electrons on the oxygen atom and the four carbon π electrons, one would expect a Hammett correlation similar to substituted benzenes. In order to test this possibility, for the base-catalyzed mechanism, several 5-substituted furfurals were prepared and oxidized under identical conditions. Figure 1 shows that I is oxidized faster with electron-withdrawing groups than with electron-releasing groups,

(37) Exchange between the carboxylate group and hydroxide ion is unlikely because of electrostatic repulsion. However, it has been found that benzoic acids will undergo slight exchange under drastic conditions, e.g., 10 days at 100°.[38]

(38) C. A. Bunton, A. E. Comyns, J. Graham, and J. R. Quayle, *J. Chem. Soc.*, 3817 (1955).

(39) N. F. Hall and O. R. Alexander, *J. Amer. Chem. Soc.*, **62**, 3455 (1940).

(40) G. A. Mills, *ibid.*, **62**, 2833 (1940).

(41) Saturated potassium permanganate solution exchanges slowly with water at 100° in neutral solution.[39] The rate of exchange appears to be only slightly faster in basic solution,[5,39,40] which is negligible for this work.

(42) S. M. Taylor and J. Halpern, *J. Amer. Chem. Soc.*, **81**, 2933 (1959).

(43) F. Freeman and A. Yeramyan, *Tetrahedron Lett.*, 4783 (1968).

(44) Curiously, attempts to prepare furfural-d_1 via the lithium aluminum deuteride reduction of furil were unsuccessful.[30] Although furfuryl alcohol was oxidized to I by manganese dioxide, the increased bond strength in α,α-d_2-furfuryl alcohol appeared to make the rate of oxidation of the aldehyde faster than the rate of oxidation of deuterated alcohol.[45]

(45) F. Freeman and A. A. Kamego, unpublished data, 1967.

(46) The ionic strength is too high to permit any detailed interpretation according to the Debye–Hückel theory.
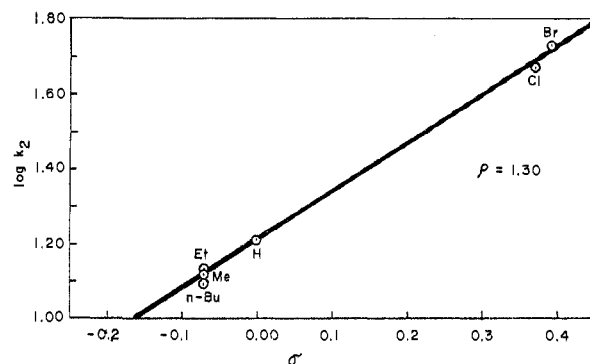
Figure 1.—Hammett plot for furfurals at pH 13.3.

Indeed, V is immeasurably fast in our stopped-flow system. Statistical treatment[47] of the rate data, using Hammett σ *meta*-substituent values, gives a ρ value of +1.30 with a correlation coefficient r of 0.9974 and a standard deviation s of 0.0235. Using σ *para*-substituent values, the statistical treatment gives a ρ of +1.52 with r = 0.9766 and s = 0.0706. The excellent correlation with σ-*meta* values is surprising, since molecular orbital calculations[48] and dipole moment measurements[49] that the *ortho*, *meta*, and *para* positions of benzene corresponds to the 3, 4, and 5 positions in 2-substituted thiophenes. Also, Noyce and Kaiser[16] obtained an excellent correlation (0.99) by using σ$^+$ *para*-substituent constants for the 5 position and σ$^+$ *meta*-substituent constants for the 4 position in the solvolysis of furylmethylcarbinol derivatives. However, it has been noted that 5-substituted 3-furoic acids correlate equally well with σ *meta*- and σ *para*-substituent constants.[17,50]

Although the observed kinetic data support Scheme II and are consistent with Scheme I, additional kinetic studies are required to fully elucidate the mechanisms in the neutral and acid regions. It is also of interest to note that the kinetics of permanganate ion oxidation of I are remarkably similar to the alkaline permanganate ion oxidation of benzaldehyde.[5,50-52]

Registry No.—I, 98-01-1; III, 1899-24-7; IV, 620-02-0; V, 698-63-5; 5-ethylfurfural, 23074-10-4; 5-chlorofurfural, 21508-19-0; 5-isobutylfurfural, 23115-88-0; 2,4-dinitrophenylhydrazone of 5-isobutylfurfural, 23074-12-6; 5-n-butylfurfural, 23074-13-7; 2,4-dinitrophenylhydrazone of 5-n-butylfurfural, 23074-14-8.

(47) (a) H. H. Jaffe, *Chem. Rev.*, **53**, 191 (1953). (b) One would expect the rate-determining step to have a negative ρ value, since electron-releasing groups should facilitate a hydride ion transfer (Scheme II). Since the observed rates depend on the equilibrium concentration of VI, it appears that ρ$_{eq}$ in the equation

$$\log(k/k_0)_{obsd} = \log (K/K_0)_{eq} + \log (k/k_0)_{rate} = \sigma(\rho_{eq} + \rho_{rate})$$

is opposite in sign and larger in magnitude than ρ$_{rate}$. This has also been observed in the permanganate oxidation of benzaldehyde and in the chromic acid oxidation of alcohols.[5]

(48) L. Melander, *Ark. Kemi*, **11**, 397 (1957).

(49) R. Keswani and H. Freiser, *J. Amer. Chem. Soc.*, **71**, 1789 (1949).

(50) On the other hand, since the measured relative substituent effect (alkyl and halogen) arises from polar effects and electrostatic interactions, the excellent correlation with σ$_{meta}$ values is not unexpected. When the data (Me, Et, H, Br, Cl) from the oxidation are correlated with the σ normal (σ$_{para}$) parameter[17b] and the σ$_{meta}$°, and σ$_{para}$° parameters (Me, H, Br, Cl),[17c] the ρ values, r, and s are 1.54, 0.988, and 0.013; 1.29, 0.998, and 0.012; and 1.47, 0.978, and 0.080, respectively.

(51) Early work in the benzaldehyde-permanganate ion reaction indicated a fractional dependence on hydroxyl ion concentration.[5] However, more recent studies have shown that this is not the case.[52]

(52) K. B. Wiberg and F. Freeman, unpublished results.