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Mechanism of oxidation of reducing sugars by osmium tetraoxide in alkaline medium by the stopped-flow technique

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ABSTRACT

The kinetics of reduction of osmium tetraoxide by L-sorbose, D-galactose, D-xylose, and L-arabinose in aqueous alkaline medium exhibit a pseudo-unimolecular rate with respect to [osmium tetraoxide]. First-order kinetics were observed with respect to both [sugar] and [OH^-] at their lower concentrations, but tends towards zero order with respect to each at higher concentrations. The experimental results suggest formation of an activated complex between enediol and osmium tetraoxide which slowly decomposes into Os(VI) species and intermediate products. The reaction products were identified, and results were also confirmed by statistical regression-analysis data.

INTRODUCTION

The kinetics of oxidative degradation of sugars by metal complexes have been the subject of numerous investigations[1–8]. Oxidants used include $K_3Fe(CN)_6$ (ref. 1), Cu(II) (ref. 2), Ce(IV) (ref. 3), V(V) (ref. 4), Cr(VI) (ref. 5), H_2O_2 (ref. 6), Ru(III)-catalysed *N*-bromosuccinimide[7], and pyridinium fluorochromate[8]. Recently the kinetics of reduction of dodecatungstocobaltate(III) by D-fructose, D-glucose, and D-mannose in aqueous medium have been investigated[9].

This paper describes experiments designed to examine pathways for the oxidation of L-sorbose, D-galactose, D-xylose, and L-arabinose by osmium tetraoxide in aqueous alkaline medium by use of the stopped-flow technique.

STOICHIOMETRY AND PRODUCTS

Several reaction mixtures containing [$OsO_4(OH)_2^{2-}$] > [sugar] at fixed [OH^-] were prepared. After the completion of reaction, the [Os(VI)] produced was determined spectrophotometrically. In all instances the reaction stoichiometry was found to be 8:1 osmium tetraoxide:sugar, and thus the reactions may all be represented as follows: $C_6H_{12}O_6$ + 8 Os(VIII) → 8 Os(VI) + other products.

The organic products were identified by paper and thin-layer chromatography[10].

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The final oxidation products from the sugars were characterised as formic and glyceraric ("tartronic") acids.

Materials. — Stock solutions of sugars (E. Merck), NaOH and KCl A.R. (B.D.H.) were prepared by direct weighing. A solution of OsO_4 (Johnson Matthey & Co., Ltd.) was prepared by dissolving the sample in an NaOH solution of known strength. Purified N_2 was passed through each solution in order to remove dissolved O_2.

Kinetics. — Progress of the reactions was measured by observing the disappearance of OsO_4 at 420 nm and the rate-constant determinations were made by using a Hi-Tech-SF-3A stopped-flow spectrometer equipped with a computerised data-acquisition and processing unit. The temperature in the stopped-flow cell was maintained with a solid-state control-unit with heaters and circulator. First-order plots of the data together with the standard deviations and the values of the first-order rate constants were obtained. The rate constants given in the tables are the averages of two or more determinations.

RESULTS AND DISCUSSION

Through a fifty-fold variation in sugar concentration, pseudo-first-order kinetics was observed with respect to sugars at lower concentrations and tends toward zero order at higher concentrations (Fig. 1).

Examination of Table I reveals that the first-order rate constants, (k_1 values) are practically constant suggesting first-order kinetics with respect to osmium tetraoxide.

The reaction follows first-order kinetics with respect to hydroxide ion at lower concentrations and tends toward zero order towards higher $[OH^-]$ (Fig. 2).

Although change in the ionic strength of the medium did not affect the reaction velocity, all experiments were nevertheless carried out at constant ionic strength.

On the basis of the foregoing experimental results, the following rate law might be proposed at low concentration of sugars and hydroxide ion:

$$\frac{-d[OsO_4(OH)_2^{2-}]}{dt} = k_1 [S] [OH^-] [OsO_4 (OH)_2^{2-}] \qquad (1)$$

where [S] denotes the concentration of sugar.

Osmium tetraoxide in alkaline medium has been reported to exist as octahedral complexes of the form[11] *trans*-$OsO_4 (OH) (H_2O)^{1-}$ and $OsO_4 (OH)_2^{2-}$, according to the following equilibrium: $OsO_4 (OH)(H_2O)^{1-} + OH^- \rightleftharpoons OsO_4 (OH)_2^{2-} + H_2O$.

It has also been reported that $OsO_4 (OH)_2^{2-}$ is the only reactive species of osmium tetraoxide in alkaline medium and that the equilibrium lies mainly towards the right[12,13].

On the basis of the current experimental results, the following scheme for the oxidation of sugars in alkaline medium may be proposed.

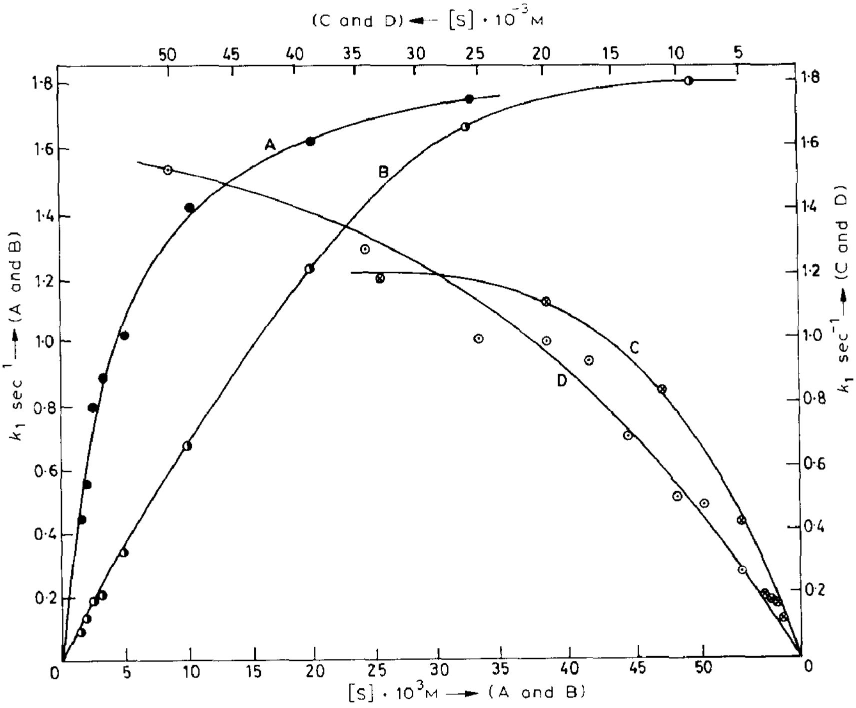


Fig. 1. Effect of variation of sugar concentration on reaction rate; plot of k_1 vs. [S]; ●, A, sorbose; ◑, B, galactose; ⊙, D, xylose; $[OsO_4(OH)_2{}^{2-}]$ 1.17×10^{-3}M; [NaOH] 4.23×10^{-1}M; [KCl] 0.1M; μ 5.22×10^{-1}M; Temp. = 35°. ⊗, C, arabinose; $[OsO_4(OH)_2{}^{2-}]$ 9.43×10^{-4}M [NaOH] 2.82×10^{-1}M [KCl] 7.2×10^{-2}M μ 3.65×10^{-1}M Temp. 35°.

$$S + OH^- \overset{K}{\rightleftharpoons} \underset{(\text{en})}{\text{enediol}} \qquad \text{(I)}$$

$$\text{en} + OsO_4(OH)_2{}^{2-} \overset{K_1}{\rightleftharpoons} \text{complex} \qquad \text{(II)}$$

$$\text{Complex} \overset{k}{\rightarrow} OsO_2(OH)_4^{2-} + \text{intermediate product} \qquad \text{(III)}$$

According to the foregoing scheme, the complex formed in step (II) decomposes slowly into an intermediate species. The formation of osmium(VI) species has been reported[14,15] previously. Mayell[11] has provided spectrophotometric evidence for the formation of osmium(VI) species in the ferricyanide oxidation of alkenes in the presence of osmium(VIII) used as catalyst.

Thus, from the foregoing mechanistic steps, the total concentration of osmium (VIII) may be given as:

$$[Os(VIII)]_T = [OsO_4(OH)_2{}^{2-}] + [\text{complex}] \qquad (2)$$

With the help of these equations, the final rate in terms of decreasing concentration of osmium(VIII) may be given as:

TABLE I

Effect on the reaction rate of varying the osmium tetraoxide concentration[a]

$[OsO_4(OH)_2^{2-}] \times 10^4$M	k_1, sec^{-1}			
	Xylose	*Arabinose*	*Galactose*	*Sorbose*
1.96			0.95	
3.14	1.29	1.23	1.66	1.02
4.71	0.99	1.37	1.24	0.97
6.29	1.00	1.14	1.14	1.16
7.86	1.26	1.67	1.53	1.65

[a] [sugar] = 3.33×10^{-2}M; [NaOH] = 4.09×10 M; [KCl] = 0.1M; μ = 5.11×10^{-1}M; temp. = 35°.

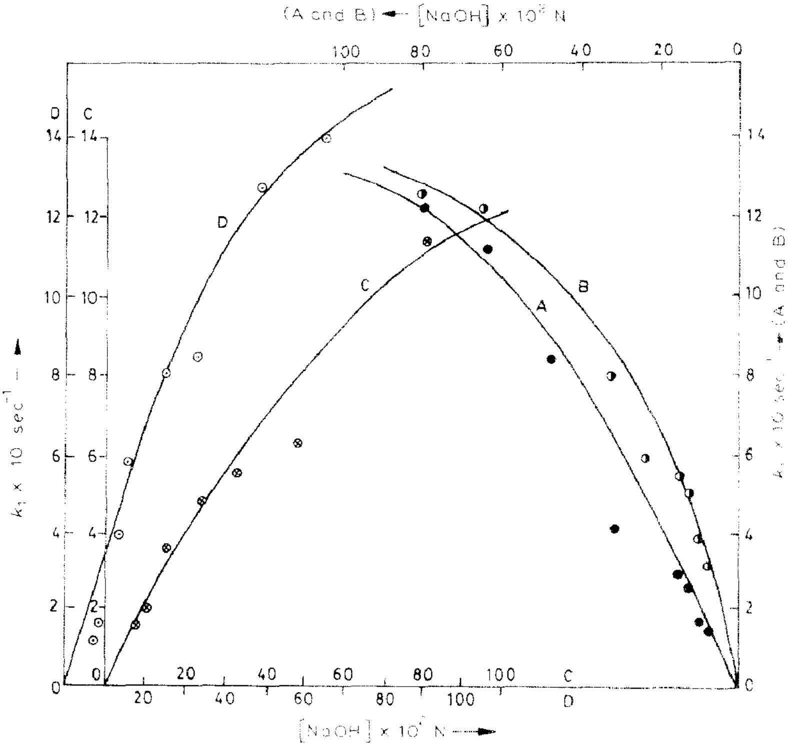


Fig. 2. Effect of variation of hydroxide ion concentration on reaction rate; plot of k_1 *vs.* $[OH^-]$; ●, A, [sorbose] = 3.33×10^{-3}M; ◑, B, [galactose] 3.33×10^{-2}M; ⊗, C, [arabinose] 1.0×10^{-2}; ⊙, D, [xyose] 3.33×10^{-2}M; $[OsO_4(OH)_2^{2-}]$ 7.86×10^{-4}M; μ = 1.66M; temp. = 35°.

$$\frac{-\mathrm{d}[\mathrm{OsO_4(OH)_2^{2-}}]}{\mathrm{d}t} = \frac{2kKK_1\,[\mathrm{S}]\,[\mathrm{OH^-}]\,[\mathrm{Os(VIII)}]_\mathrm{T}}{1 + KK_1\,[\mathrm{S}]\,[\mathrm{OH^-}]} \tag{3}$$

and

$$\frac{-\mathrm{d}[\mathrm{OsO_4(OH)_2^{2-}}]}{\mathrm{d}t \times [\mathrm{Os(VIII)}]_\mathrm{T}} = k_1 = \frac{2kKK_1\,[\mathrm{S}]\,[\mathrm{OH^-}]}{1 + KK_1[\mathrm{S}]\,[\mathrm{OH^-}]} \tag{4}$$

The rate laws (*3*) and (*4*) clearly explain the first-order kinetics with respect to osmium tetraoxide concentration. The retarding effect of [S] as well as of [OH⁻] on reaction rate may also be explained from Eqs. (*3*) and (*4*).

At low concentrations of the substrate and hydroxide ion, it might be safely assumed that the inequality $1 >> KK_1\,[\mathrm{S}]\,[\mathrm{OH^-}]$ under such conditions is valid and Eq. (*4*) reduces to:

$$k_1 = 2kKK_1\,[\mathrm{S}]\,[\mathrm{OH^-}] \tag{5}$$

Eq. (*5*) explains the observed first-order kinetics at low concentrations of sugar and of hydroxide ion. The calculated kKK_1 values are been given in Table II. The close similarity in kKK_1 values obtained by two different methods [namely (*i*) from a plot of k_1 *vs.* [sugar], Fig. 1; and (*ii*) from a plot of k_1 *vs.* [OH⁻]; Fig. 2] provides support for the rate-law equation (*5*).

At sufficienly high concentrations of hydroxide ion and sugars, the inequality $KK_1\,[\mathrm{S}]\,[\mathrm{OH^-}[\; > > 1$ will be valid. Under such conditions the rate-law equation (*4*) reduces to:

$$k_1 = \frac{-\mathrm{d}[\mathrm{Os(VIII)(OH)_2^{2-}}]}{[\mathrm{Os(VIII)}]_T \times \mathrm{d}t} = 2k \tag{6}$$

Eq. (*6*) explains the observed zero-order kinetics at higher concentrations of hydroxide ion and of sugar.

In order to verify the rate-law equation (*4*) for concentrations of hydroxide ion and of sugars in an intermediate range, Eq. (*4*) might be written in the form:

$$\frac{1}{k_1} = \frac{1}{2kKK_1[\mathrm{S}][\mathrm{OH^-}]} + \frac{1}{2k} \tag{7}$$

Eq. (*7*) predicts that the plots of the reciprocal of the velocity constant (namely, $1/k_1$) versus the reciprocal of [sugar] at fixed [OH⁻], or of [OH⁻] at fixed [sugar], should be straight lines with positive intercepts at the Y-axis; such plots were actually observed (Figs. 3 and 4). The values of kKK_1 and k obtained from the slope and intercept respectively of these straight lines are recorded in Tables II and III, respectively.

Inspection of Tables II and III reveals that the values of kKK_1 and k obtained from different plots, based on Eqs. (*5*) and (*7*) are in good agreement with one another. These results support the validity of the rate-law equation (*4*) and thereby of the proposed reaction mechanism.

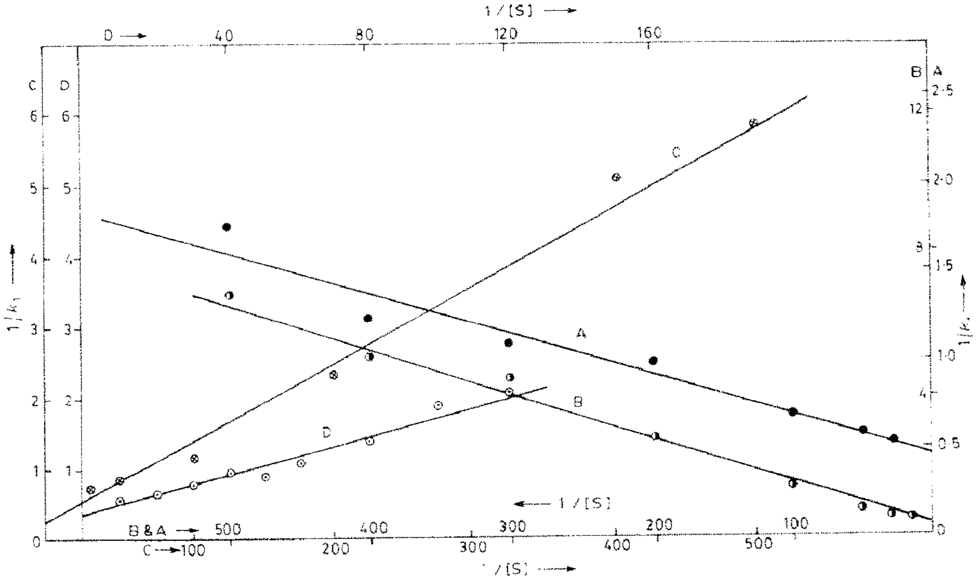


Fig. 3. Plot of $1/k_1$ *vs.* $1/[S]$. ⊙, D, xylose; ⊗, C, arabinose; ◑, B, galactose; ●, A, sorbose (note: captions are same as in Fig. 1).

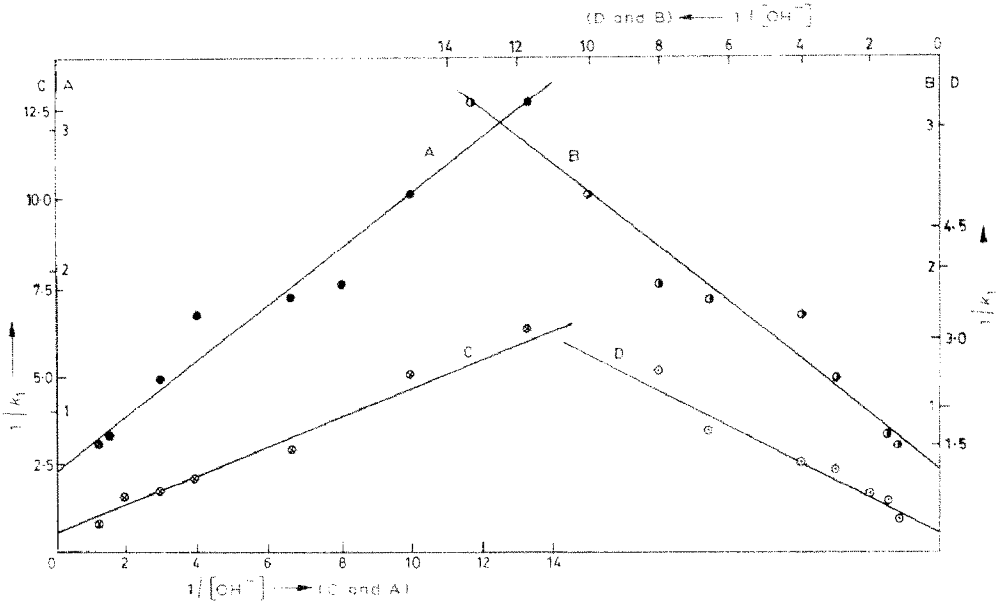


Fig. 4. Plot of $1/k_1$ *vs.* $1/[OH^-]$; ⊗, C, arabinose; ⊙, D, xylose; ●, A, sorbose; ◑, B, galactose (note: captions are same as in Fig. 2)

observed (Figs. 3 and 4). The values of kKK_1 and k obtained from the slope and intercept respectively of these straight lines are recorded in Tables II and III, respectively.

Inspection of Tables II and III reveals that the values of kKK_1 and k obtained from different plots, based on Eqs. (*5*) and (*7*) are in good agreement with one another. These results support the validity of the rate-law equation (*4*) and thereby of the proposed reaction mechanism.

TABLE II

Values of kKK_1 obtained from plots showing variations of k_1 *vs.* [sugar]; $1/k_1$ *vs.* 1/[sugar]; k_1 *vs* [OH]; and $1/k_1$ *vs.* $1/[OH^-]$

Sugar	$kKK_1 \times 10^{-2}$			
	From k_1 vs. *[S]*	*From* $1/k_1$ vs. *1/[S]*	*From* k_1 vs. $[OH^-]$	*From* $1/k_1$ vs. $1/[OH^-]$
Xylose	1.42	1.73	1.20	1.20
Arabinose	3.40	3.19	2.50	2.50
Galactose	1.67	1.82	1.35	1.60
Sorbose	7.48	7.95	6.67	6.68

TABLE III

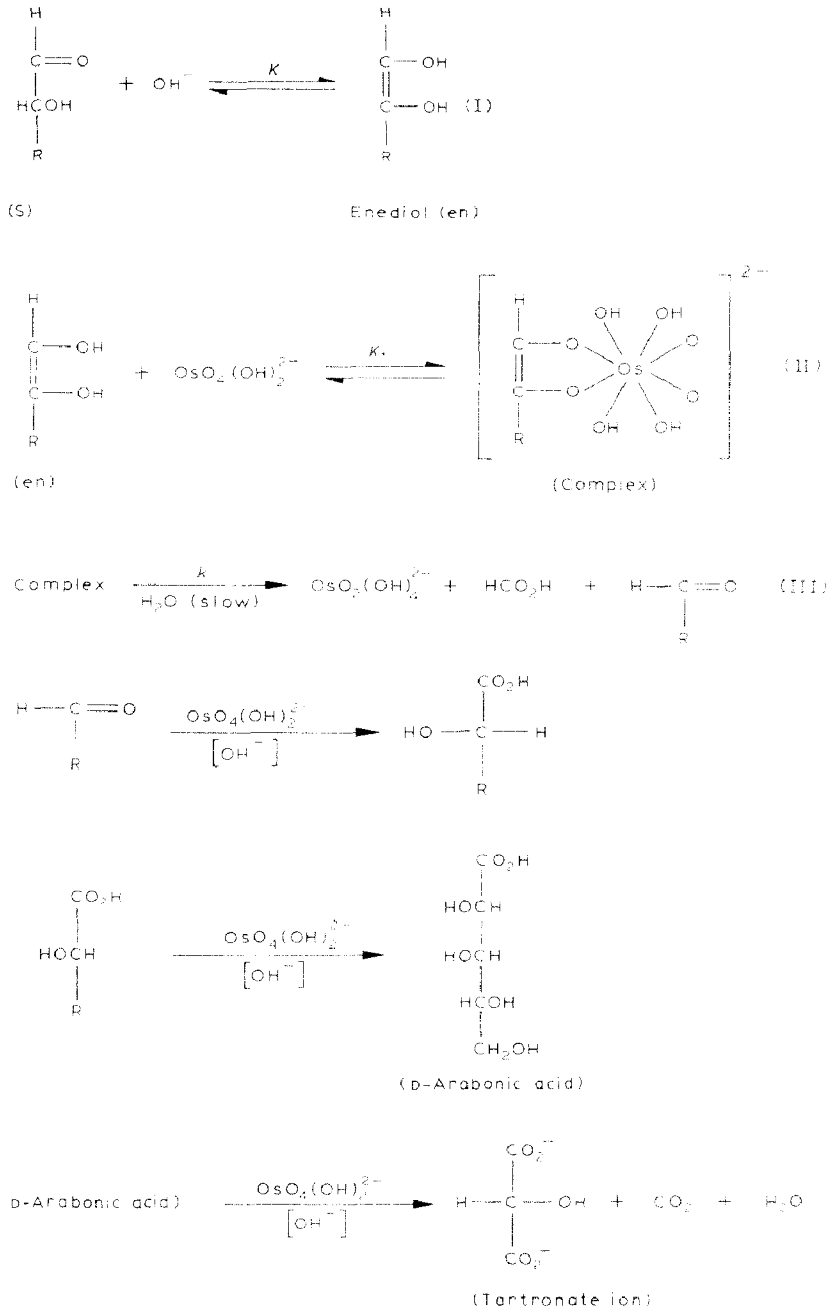
Values of k obtained from plots, showing variations of $1/k_1$ *vs.* 1/[sugar] and $1/k_1$ *vs.* $1/[OH^-]$

Sugar	k sec^{-1}		
	From $1/k_1$ vs. *1/[S]*	*From* $1/k_1$ vs. $1/[OH^-]$	*Average value*
Xylose	1.43	1.66	1.54
Arabinose	1.43	1.00	1.22
Galactose	1.00	0.90	0.95
Sorbose	1.00	1.00	1.00

In the oxidation of reducing sugars by alkaline ferricyanide[1] and Cu(II)[2], Singh *et al.* have shown that the reaction occurs *via* the formation of an enediol. They have further shown that the rate of enolization is the same as the rate of oxidation. However, in the present case, the rate of disproportionation of the activated complex is the lowest one and is hence the rate-determining step.

Such complexes have also been indicated between osmium tetraoxide and glycols[16], alcohols[16], diols[16], ketones[14], and hydroxy acids[15]. In alkaline medium, the reaction occurs *via* the formation of an activated complex between permanganate ion and diols[17].

In the oxidation of the foregoing sugars, chromatography reveals the formation of glycerarate (tartronate) ion as the final oxidation-product. On the basis of these results, the full sequence of the steps may be given as follows:



STATISTICAL REGRESSION-ANALYSIS DATA

To assess the probability of the proposed reaction mechanism, the experimental data were subjected to statistical regression analysis. Theoretical values were obtained by using equations developed from the proposed reaction mechanism. The results

TABLE IV

Statistical regression-analysis data

Sugars	*Correlation coefficient*	*Ratio of variance*	*Standard error of estimation*
Xylose	0.98	171.5	0.24
Arabinose	0.97	99.1	0.74
Galactose	0.98	212.9	0.64
Sorbose	0.96	85.4	0.16

obtained for the various sugars are given in Table IV, and indicate that: (*a*) the correlation coefficients for the different sugars are very close to unity, (*b*) the ratio of variance obtained for the different sugars is quite high; and (*c*) the standard errors of estimation are quite low for each of the sugars studied.

All the foregoing facts provide reasonably good support for the proposed mechanism.

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REFERENCES

1 N. Nath and M. P. Singh, *J. Phys. Chem.*, 69 (1965) 2038–2043.
2 S. V. Singh, O. C. Saxena, *J. Am. Chem. Soc.*, 92 (1972) 537–541.
3 R. N. Mehrotra and E. S. Amis, *J. Org. Chem.*, 39 (1974) 1788–1791.
4 A. Kumar and R. N. Mehrotra, *J. Org. Chem.*, 40 (1975) 1248–1252.
5 K. K. Sengupta and S. N. Basu, *Carbohydr. Res.*, 86 (1980) 7–16.
6 H. S. Isbell and H. L. Frush, *Carbohydr. Res.*, 28 (1973) 295–301.
7 T. Kistayya, M. S. Reddy, and S. Kandlikar, *Indian J. Chem. Sect. A*, 25 (1986) 905–907.
8 R. Varadarajan and R. K. Dhar, *Indian J. Chem., Sect A*, 25 (1986) 474–475.
9 Mala Gupta, Swapan K. Saha, and Pradyot Banerjee, *J. Chem. Soc. Perkins Trans.*, 2 (1988) 1781–1785.
10 R. D. Hartly and G. J. Lawson, *J. Chromatogr.*, 4 (1960) 410–413.
11 J. S. Mayell, *Ind. Eng. Chem.*, 7 (1968) 129–136.
12 (*a*) W. P. Griffith, *Q. Rev. Chem. Soc.*, 19 (1965) 254–273; (*b*) F. A. Cotton and G. Wilkinson, *Advanced Inorganic Chemistry*, 2nd ed., Wiley Eastern Pvt. Ltd., New Delhi (1966) pp. 1007 and 993.
13 W. P. Griffith, *The Chemistry of Rare Platinum Metals*, Interscience, London, New York, (1967) pp. 42–125.
14 N. P. Singh, V. N. Singh, H. S. Singh, and M. P. Singh, *Aust. J. Chem.*, 21 (1968) 2913–2918; 23 (1970) 921–928.
15 V. N. Singh, H. S. Singh, and B. B. L. Saxena, *J. Am. Chem. Soc.*, 91 (1969) 2643–2648.
16 (a) H. S. Singh, S. P. Singh, S. M. Singh, R. K. Singh, and A. K. Sisodiya, *J. Phys. Chem.*, 79 (1975) 1920–1924; (b) H. S. Singh, V. P. Singh, J. M. Singh, and P. N. Srivastava, *J. Catal.* 49 (1977) 135–140; (c) H. S. Singh, A. K. Sisodiya, S. M. Singh, R. K. Singh, and R. N. Singh, *J. Chim. Phys.*, 3 (1976) 283–286.
17 K. B. Wiberg, *Oxidation in Organic Chemistry*, Academic Press, New York, Part A, (1965) pp. 64.