Reactivities of D-Mannitol and Related Alcohols toward Chromium(VI) and Chromium(IV)

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The kinetics of the reactions between Cr(VI) and a homologous series of reductants, formed by CH_3OH and polyalcohols of the type $CH_2OH(CHOH)_{n-2}CH_2OH$ with $n = 2-6$, have been studied in perchloric acid aqueous solutions. The rate law was: $v = k_4[\mathrm{Cr(VI)}][\mathrm{alcohol}][\mathrm{H}^+]^2$. The activation parameters were in the range $\Delta H^\circ_{\neq} = 37.4-43.1$ kJ mol^{-1} and $\Delta S^\circ_{\neq} = (-168) - (-139)$ J K^{-1} mol^{-1}. The reactions were inhibited by the Cr(IV)-trapping agent Mn(II), and this effect was used to obtain the rate constants for the 2-electron reduction of the intermediate Cr(IV) by the alcohols ($k_2 = 52.2-166$ M^{-1} s^{-1}). A mechanism is proposed according to which Cr(VI) is reduced by the alcohol of n carbon atoms through $n(n+1)/2$ different reaction pathways involving chromate monoesters and diesters as intermediates, whereas Cr(IV) is reduced through n reaction pathways by direct hydride-ion transfer.

Introduction

Recently, Espenson and co-workers[1,2] have revised the general mechanism for the reactions of chromium(VI) with many organic substrates in acidic aqueous solutions, showing that, although discarded for a long time by most chemists interested in those reactions, one of the original mechanisms proposed half a century ago by Westheimer[3] is actually the predominant reaction pathway. In that mechanism the organic substrate behaves toward both the reactant Cr(VI) and the intermediate Cr(IV) as a 2-electron reductant. We confirmed by an independent method the validity of that mechanism for the Cr(VI) oxidations of formic acid,[4,5] formaldehyde,[5,6] and several primary and secondary monoalcohols.[5] Now we have confirmed that the same mechanism applies to a family of polyalcohols.

On the other hand, although the reactions of Cr(VI) with most organic substrates are inhibited by Mn(II),[7] for some polyhydroxylic compounds (including α-hydroxyacids,[8] dicarboxylic acids,[8,9] and the saccharide D-ribose[10]) catalysis by Mn(II) has been found. However, for the Cr(VI) oxidation of the family of polyalcohols studied in the present work Mn(II) acts in its usual role as an inhibitor, thus suggesting that its role as catalyst is specific of those polyhydroxylic organic reductants capable of stabilizing Mn(III) as an intermediate.

We were especially interested in the reaction between Cr(VI) and mannitol because the latter has been shown to suppress the Cr(VI) toxicity in vivo toward a catalase-defective bacterial strain, and it has been suggested that the effect of mannitol is due to its capability of acting as a hydroxyl radical scavenger, rather than to a direct interaction between Cr(VI) and mannitol.[11] According to this hypothesis, hydroxyl free radicals generated in the intracellular metabolization of Cr(VI) (probably involving biologically important substances such as ascorbic acid,[12] glutathione,[13] or hydrogen peroxide[14]) might play an important role in the expression of chromate-induced toxicity and carcinogenesis.[15] The results found in the present work indicate that Cr(VI) and mannitol do not interact appreciably at physiological pH, thus supporting the free radical hypothesis.

Experimental Section

Materials and Methods. The solvent was water previously purified by deionization, distillation, and circulation through a Millipore system. The oxidant in our media was predominantly in the form of monomeric Cr(VI) ($HCrO_4^-$), obtained by dissolution of $K_2Cr_2O_7$. The six reductants studied belonged to a homologous series of alcohols formed by CH_3OH (methanol) and polyalcohols of the type $CH_2OH(CHOH)_{n-2}CH_2OH$ with $n = 2$ (ethylene glycol), 3 (glycerol), 4 (DL-threitol), 5 (adonitol) and 6 (D-mannitol), and their aqueous solutions were prepared right before their use. The acidity of the solutions was controlled by means of $HClO_4$ in concentration low enough (≤ 0.395 M) to be almost completely dissociated. The ionic strength was controlled by the use of $NaClO_4$. The inhibitor used was $MnSO_4$. All the inorganic substances used were obtained from Merck and the organic ones from Sigma.

The reactions were monitored in thermostated quartz cuvettes (path length 1 cm) at one of the absorption maxima corresponding to $HCrO_4^-$ (352 nm) using a Varian Cary 219 UV−Vis spectrophotometer. The kinetic study was complete for the Cr(VI) oxidation of D-mannitol, whereas for the oxidation of the other alcohols only those aspects for which a significant difference with the behavior of D-mannitol was expected [stoichiometry, inhibition by Mn(II), and activation parameters] were studied.

Stoichiometric Experiments. For each alcohol 20 reacting mixtures with $[HCrO_4^-]_o$ fixed and $[\mathrm{alcohol}]_o$ variable were prepared and introduced in hermetically closed vials. Due to the extreme slowness of the reactions under dilute reductant conditions, they were allowed to progress in the darkness and at room temperature for at least 378 days, and then the final absorbances of the solutions at 352 nm were measured.

Kinetic Experiments. Each kinetic run was followed until 25% of the reaction was completed. The absorbance-time data were fitted to a function of the type

$$\ln(A_t - A_\infty) = a_o + a_1 t + a_2 t^2 \tag{1}$$

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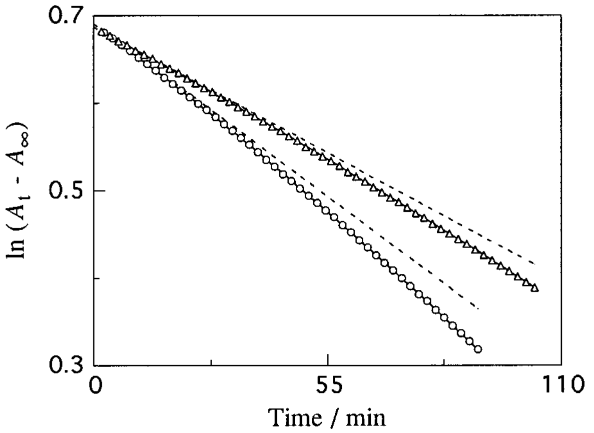


Figure 1. Attempted pseudo first-order plots for the reaction of Cr(VI) (1.28×10^{-3} M) with D-mannitol (6.65×10^{-3} M) in the presence of $HClO_4$ (0.366 M) at 25.0 °C, and at $[MnSO_4] = 0$ (circles) and 2.40×10^{-4} M (triangles). The solid lines are the fits of the experimental data to a function of the type given in eq 1, whereas the dashed lines represent the tangents to the experimental curves at time = 0.

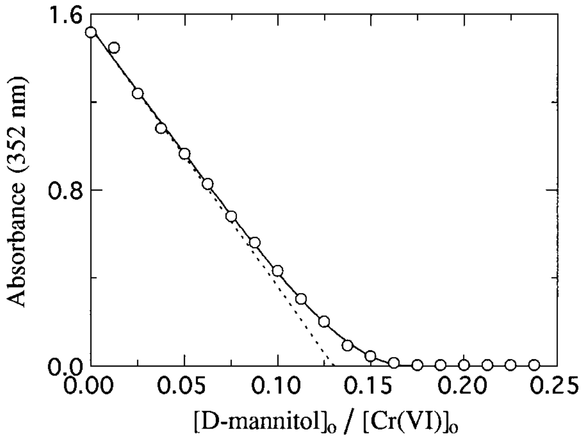


Figure 2. Dependence of the final absorbance at 352 nm on the reductant/oxidant initial concentrations ratio for the reaction of Cr(VI) (1.02×10^{-3} M) with D-mannitol in the presence of $HClO_4$ (0.366 M) at room temperature. The dashed line represents the tangent to the curve at [D-mannitol]$_o$ = 0.

and the initial rate was obtained as $v_o = -a_1[Cr(VI)]_o$. Two typical examples of the fits used to obtain the initial rates are shown in Figure 1. All the experiments were duplicated (850 kinetic experiments in total), and the kinetic data appearing in this work are the averages of the two determinations. The typical standard deviation of the initial rate was ±0.4%.

Results

Stoichiometry. The final product of the reduction of Cr(VI) by the six different alcohols was Cr(III), identified by its two weak bands in the visible region of the spectrum ($\lambda_{max,1} = 412 \pm 3$ nm, $\epsilon_1 = 20.3 \pm 0.8$ M^{-1} cm^{-1}, and $\lambda_{max,2} = 572 \pm 4$ nm, $\epsilon_2 = 17.3 \pm 0.9$ M^{-1} cm^{-1}).[14] In Figure 2, the final absorbances at 352 nm corresponding to different reductant/oxidant initial ratios for the reduction of a fixed concentration of Cr(VI) by a variable concentration of D-mannitol are shown. Given the mild curvature of the plot, indicating that even after more than one year from the beginning the reaction was not totally completed, the tangent to the curve at [reductant]$_o$ = 0 was drawn, and, from the absolute value of its slope, the number of electrons transferred from each molecule of organic substrate to Cr(VI), under conditions of excess oxidant with respect to reductant, was calculated as

$$\text{number } (e^-) = \frac{3|\text{slope}|}{A(Cr^{VI}) - A(Cr^{III})} \qquad (2)$$

where $A(Cr^{VI})$ and $A(Cr^{III})$ are the absorbances at 352 nm corresponding to Cr(VI) and Cr(III), both at the same concentration (1.02×10^{-3} M). The same procedure was followed for the reactions of Cr(VI) with the other alcohols and the results appear in Table 1.

TABLE 1: Dependence of the Number of Electrons Transferred to Cr(VI) on the Number of Carbon Atoms per Molecule of Organic Substrate[a]

carbon atoms/ molecule	number $(e^-)_{exp}$[b]	number $(e^-)_{max}$[c]
1	5.7	6
2	8.6	10
3	12.0	14
4	16.5	18
5	19.8	22
6	23.3	26

[a] $[Cr(VI)]_o = 1.02 \times 10^{-3}$ M, $[HClO_4] = 0.366$ M, room temperature. [b] Experimental values. [c] Maximum possible values corresponding to the oxidation processes: $C_nH_{2n+2}O_n + nH_2O \rightarrow nCO_2 + (4n + 2)H^+ + (4n + 2)e^-$.

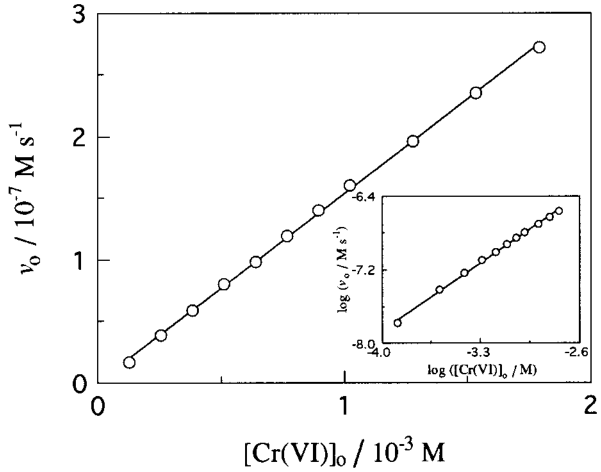


Figure 3. Dependence of the initial rate on the initial concentration of Cr(VI) for its reduction by D-mannitol (1.66×10^{-2} M) in the presence of $HClO_4$ (0.366 M) at 25.0 °C. Inset: Double-logarithmic plot of v_o vs $[Cr(VI)]_o$; slope = 1.05 ± 0.02.

Kinetics. As demonstrated by the dependence of the initial rate on the initial oxidant concentration (Figure 3), the oxidation of D-mannitol was of first order in Cr(VI). However, the pseudo first-order plots (under a large excess of reductant) showed a definite curvature (Figure 1). The ratio of the first half-life over the second half-life for some experiments was calculated as $t_{1/2}/(t_{3/4}-t_{1/2})$, where $t_{1/2}$ and $t_{3/4}$ are the time intervals required for the destruction of 50% and 75% of the initial Cr(VI). Although for a perfect pseudo first-order reaction that ratio is unity, we found ratios higher than unity. This means that the reaction showed a self-catalytic behavior, since the rate decreased more slowly than the Cr(VI) concentration did. The ratio values approached the limit expected for a pseudo first-order reaction as the reductant/oxidant initial concentrations ratio increased (Figure 4), whereas addition of Mn(II) resulted in an increase of the ratio values (Figure 4, inset). Some degree of self-catalysis was observed for the reactions of Cr(VI) with the six alcohols studied.

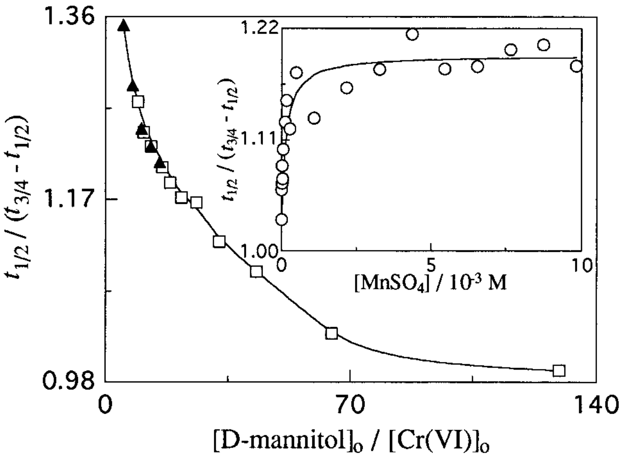


Figure 4. Dependence of the first half-life/second half-life ratio on the reductant/oxidant initial concentrations ratio at [HClO$_4$] = 0.366 M and 25.0 °C. Squares: [Cr(VI)]$_o$ = variable, [D-mannitol]$_o$ = 1.66 × 10^{-2} M. Triangles: [Cr(VI)]$_o$ = 1.28 × 10^{-3} M, [D-mannitol]$_o$ = variable. Inset: Dependence of the first half-life/second half-life ratio on the concentration of MnSO$_4$ at [Cr(VI)]$_o$ = 2.56 × 10^{-4} M and [D-mannitol]$_o$ = 1.66 × 10^{-2} M.

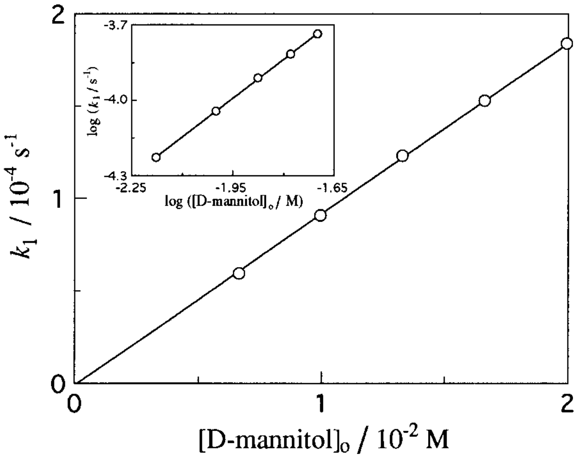


Figure 5. Dependence of the first-order rate constant on the initial concentration of D-mannitol for its oxidation by Cr(VI) (1.28 × 10^{-3} M) in the presence of HClO$_4$ (0.366 M) at 25.0 °C. Inset: Double-logarithmic plot of k_1 vs [D-mannitol]$_o$; slope = 1.03 ± 0.01.

The rate law found (supported by the data shown in Figures 3, 5,and 6) was

$$\nu_o = -\left(\frac{d[Cr(VI)]}{dt}\right)_{t=0} = k_4[Cr(VI)]_o[alcohol]_o[H^+]_o^2 \tag{3}$$

where k_4 is the experimental fourth-order rate constant, and for which a decrease with increasing ionic strength was observed (Figure 7).

The reactions of Cr(VI) with the six alcohols were inhibited by Mn(II) (Figure 8), their first-order rate constant (defined as $k_1 = \nu_o/[Cr(VI)]_o$) depending on the concentration of inhibitor according to the law

$$k_1 = \frac{a}{1 + b[Mn(II)]} + c \tag{4}$$

where b, the parameter measuring the efficiency of the inhibitor, was roughly independent of the initial concentration of oxidant but inversely proportional to that of reductant (Figure 8, inset):

$$k_1 = \frac{a}{1 + b' \dfrac{[Mn(II)]}{[alcohol]}} + c \tag{5}$$

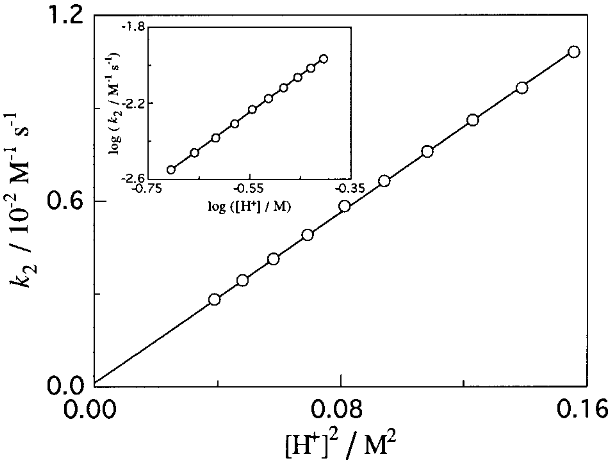


Figure 6. Effect of the hydrogen ion concentration on the second-order rate constant ($k_2 = k_1$/[D-mannitol]$_o$) for the reduction of Cr(VI) (7.68 × 10^{-4} M) by D-mannitol (1.66 × 10^{-2} M) at ionic strength 0.395 M (NaClO$_4$) and 25.0 °C. Inset: Double-logarithmic plot of k_2 vs [H$^+$]; slope = 1.95 ± 0.01.

The fourth-order rate constants for the six reactions studied fulfilled both the Arrhenius (Figure 9) and Eyring (Figure 9, right-hand inset) equations. The activation parameters determined for the different alcohols appear in Table 2. For the reaction of Cr(VI) with D-mannitol, a slight decrease of parameter b' with increasing temperature was observed (Figure 9, left-hand inset), the apparent activation energy associated to that parameter being $\Delta E_a = -11 \pm 6$ kJ mol^{-1}.

Discussion

Mechanism. The experimental results found in this work are consistent with the following mechanism for the reactions studied:

$$HCrO_4^- + H^+ \overset{K_I}{\rightleftharpoons} H_2CrO_4 \tag{6}$$

$$R_1R_2CHOH + H_2CrO_4 \overset{K_{II}}{\rightleftharpoons} R_1R_2CH{-}O{-}CrO_3H + H_2O \tag{7}$$

$$R_1R_2CH{-}O{-}CrO_3H + H^+ \overset{K_{III}}{\rightleftharpoons} R_1R_2CH{-}O{-}CrO_3H_2^+ \tag{8}$$

$$R_1R_2CH{-}O{-}CrO_3H_2^+ \xrightarrow[slow]{k_{IV}} R_1R_2C{=}O + HCrO_2^+ + H_2O \tag{9}$$

$$HCrO_2^+ + H^+ \rightarrow CrO^{2+} + H_2O \tag{10}$$

$$CrO^{2+} + R_1R_2CHOH \xrightarrow{k_V} Cr^{2+} + R_1R_2C{=}O + H_2O \tag{11}$$

$$Cr^{2+} + Cr(VI) \rightarrow Cr^{3+} + Cr(V) \tag{12}$$

$$Cr(V) + R_1R_2CHOH \rightarrow Cr^{3+} + R_1R_2C{=}O \tag{13}$$

Considering the relatively low acidity used in this work ([H$^+$] ≤ 0.395 M), Cr(VI) was present in the solutions predominantly as hydrogenchromate ion. However, a small amount was present as chromic acid (eq 6), which is the active oxidizing agent, given

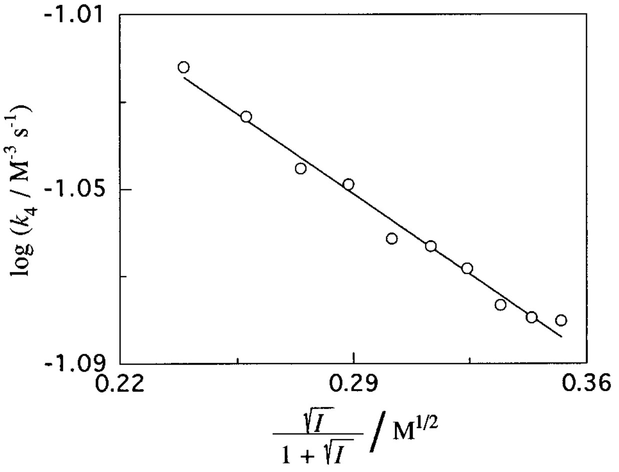


Figure 7. Effect of the ionic strength on the fourth-order rate constant for the reduction of Cr(VI) (7.68×10^{-4} M) by D-mannitol (6.65×10^{-2} M) in the presence of $HClO_4$ (9.87×10^{-2} M) at 25.0 °C. Slope = -0.53 ± 0.02 M$^{-1/2}$.

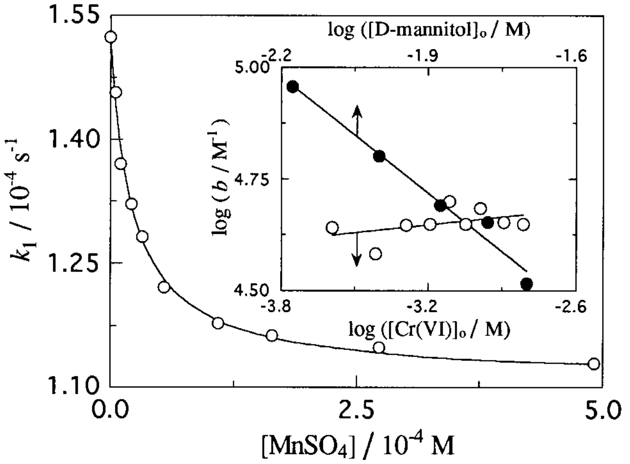


Figure 8. Effect of the concentration of $MnSO_4$ on the first-order rate constant for the reduction of Cr(VI) (1.28×10^{-3} M) by D-mannitol (1.66×10^{-2} M) in the presence of $HClO_4$ (0.366 M) at 25.0 °C. The solid line represents the fit of the experimental data to eq 4. Inset: Dependence of parameter *b* on the initial concentrations of Cr(VI) (empty circles, slope = 0.06 ± 0.04) and D-mannitol (filled circles, slope = -0.87 ± 0.08).

its low electron density compared with that of $HCrO_4^-$. In eq 7 an esterification reaction between the alcohol and chromic acid is proposed, and spectrophotometric proofs of the formation of such esters have been reported.[16,17] Protonation of a chromium-bonded oxygen atom (eq 8) destabilizes that ester by decreasing the electron density of the chromium moiety, thus enhancing the oxidizing power of chromium. As a result, the ester suffers an internal redox process in the rate-determining step (eq 9), leading to the formation of a carbonylic compound and Cr(IV). The cleavage in that step of the C−H bond of the organic substrate situated at the α position with respect to the alcohol group is consistent with the primary deuterium isotope effect reported for the oxidation of alcohols by Cr(VI) when that bond is replaced by C−D.[18] The predominant form of the intermediate Cr(IV) in acidic media, oxochromium(IV) ion,[5] is formed in a fast step (eq 10). The 2-electron reduction of the latter by the organic substrate leads to Cr(II) (eq 11), whose formation is supported by the spectrophotometric detection of superoxochromium(III) ion, $CrOO^{2+}$, when Cr(IV) is reduced in oxygenated acidic aqueous solutions by most organic substrates (with a few exceptions, such as cyclobutanol and pivaldehyde, that behave as 1-electron reductants).[2] The formation of the final inorganic product, Cr(III), occurs both by oxidation of Cr(II) by Cr(VI) (eq 12) and by reduction of Cr(V) by the organic substrate (eq 13).

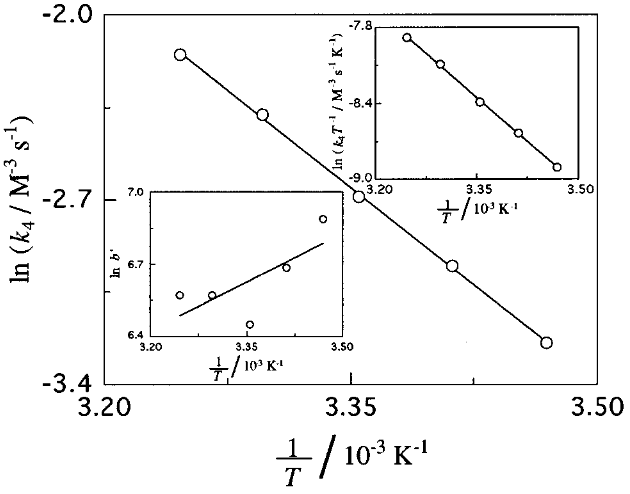


Figure 9. Arrhenius plot for the reduction of Cr(VI) (1.28×10^{-3} M) by D-mannitol (9.98×10^{-3} M) in the presence of $HClO_4$ (0.366 M) in the range 15.0−35.0 °C. Right-hand inset: Eyring plot under the same conditions. Left-hand inset: Effect of temperature on parameter *b*′ under the same conditions and [$MnSO_4$] = 0−3.61×10^{-4} M.

TABLE 2: Dependence of the Activation Parameters for the Reduction of Cr(VI) on the Number of Carbon Atoms per Molecule of Organic Substrate[a]

carbon atoms/ molecule	E_a/kJ mol^{-1}	$\Delta H^{\circ}_{\neq}$/kJ mol^{-1}	$\Delta S^{\circ}_{\neq}$/J K^{-1} mol^{-1}
1	44.6 ± 0.3	42.1 ± 0.3	-168 ± 1
2	45.6 ± 1.3	43.1 ± 1.3	-149 ± 4
3	42.6 ± 0.7	40.1 ± 0.7	-149 ± 2
4	40.8 ± 0.7	38.3 ± 0.7	-144 ± 2
5	39.9 ± 0.3	37.4 ± 0.3	-139 ± 1
6	40.8 ± 0.5	38.3 ± 0.5	-139 ± 2

[a] The activation entropies are associated to the fourth-order rate constant (in M^{-3} s^{-1}) and referred to the 1 M standard state.

The rate law deduced from the proposed mechanism is consistent with eq 3, the experimental fourth-order rate constant being

$$k_4 = 2K_{I}K_{II}K_{III}k_{IV} \tag{14}$$

For the six alcohols studied the number of electrons transferred to Cr(VI) from each molecule of organic substrate was only slightly lower than the maximum value expected for a complete oxidation (Table 1). Therefore, the main product of oxidation of the alcohols by excess Cr(VI) was CO_2. Thus, under those conditions, the carbonylic compounds (aldehydes and ketones formed, respectively, from the oxidation of terminal and nonterminal alcohol groups) generated in the initial stages of the reaction[19] must be considered as long-lived intermediates, eventually oxidized by the excess of oxidant, rather than as true reaction products. However, this conclusion cannot be applied to the conditions corresponding to the kinetic experiments, where a large excess of reductant was used in most cases.

In fact, the self-catalytic behavior observed in the absence of Mn(II) was caused by the accumulation of intermediate aldehydes, given the higher reactivity of the latter toward Cr(VI) with respect to that of alcohols. This hypothesis agrees with previous interpretations,[10,20,21] as well as with our finding that the deviation from a pseudo first-order behavior was controlled

TABLE 3: Dependencies of Parameter b' and the Rate Constants for the Reduction of Cr(VI) and Cr(IV) on the Number of Carbon Atoms per Molecule of Organic Substrate[a]

carbon atoms/ molecule	$10^{-3} b'$	k_4 (Cr^{VI})/ 10^{-4} M^{-3} s^{-1}	k_2 (Cr^{IV})/ M^{-1} s^{-1}
1	2.08	4.31	52.2[b]
2	1.53	28.9	71.1[c]
3	1.30	90.3	83.6[c]
4	0.98	383	111[c]
5	0.76	926	143[c]
6	0.65	678	166[c]

[a] [$HClO_4$] = 0.366 M, 25.0 °C. [b] Taken from ref 2. [c] Obtained from the b' values determined in this work, taking as reference the value corresponding to methanol.

by the [reductant]$_o$/[oxidant]$_o$ ratio, as illustrated in Figure 4 for two different series of experiments (one with [Cr(VI)]$_o$ constant and [D-mannitol]$_o$ variable, and the other with [Cr(VI)]$_o$ variable and [D-mannitol]$_o$ constant). The self-catalysis decreased as that ratio increased because the intermediate aldehyde competed with the reactant alcohol.

Effect of Manganese(II). The rate law found for the inhibition of the reactions by Mn(II) (eq 5) suggests that the latter and the alcohol compete for the same species. This must be the intermediate Cr(IV), given the well-known ability of Mn(II) to act as a trapping agent for that intermediate:[7]

$$CrO^{2+} + Mn^{2+} + H_2O \xrightarrow{k_{VI}} Cr^{3+} + Mn^{3+} + 2OH^- \quad (15)$$

Now, assuming that CrO^{2+} is in steady state, we obtain eq 5 with $a = c = K_I K_{II} K_{III} k_{IV}[\text{alcohol}][H^+]^2$ and $b' = k_{VI}/k_V$. The experimental values obtained for parameters a and c were of the same order of magnitude, but with $a < c$ in all cases. This means that Mn(II) was a less efficient inhibitor than expected from the mechanism proposed, and it can be easily explained if we consider that, although most of the Cr^{2+} formed in the absence of Mn(II) is oxidized by Cr(VI) (eq 12), some Cr^{2+} may also be oxidized by either Cr(IV) or dissolved O_2,[2] but only the path of disappearance of Cr^{2+} corresponding to the oxidation by Cr(VI) can be inhibited by Mn(II). The k_{VI}/k_V ratios were determined for the oxidation of the six alcohols studied and, used along with the value of k_V reported for the reaction between Cr(IV) and methanol,[2] we obtained the values of k_V for the different alcohols. Those values are given in Table 3 in the form of the experimental second-order rate constants for the reactions between Cr(IV) and the different alcohols ($k_2 = v/[\text{Cr(IV)}]\,[\text{alcohol}]$). The difference between the theoretical rate constant for these reactions, k_V, and the experimental rate constant, k_2, is that, whereas the former refers to the elementary step corresponding to a single pathway (eq 11), the latter refers to the global value obtained from the combination of the n reaction pathways involved in the oxidation of the alcohol of n carbon atoms by Cr(IV).

The value obtained for k_{VI} was 1.09×10^5 M^{-1} s^{-1}, in good agreement with the one previously reported.[5] That value, much higher than those obtained for k_V in the six reactions studied, indicates that the reduction of CrO^{2+} by Mn^{2+} (electron transfer, eq 15), despite involving the encounter between two like-charged ionic reactants, is more favorable than the reductions of CrO^{2+} by the organic substrates (hydride-ion transfer, eq 11). This was confirmed by the finding that, in the case of D-mannitol, from the dependence of parameter b' on temperature (Figure 9, left-hand inset) it could be inferred that the activation energy for the reduction of Cr(IV) by D-mannitol was slightly higher than that for its reduction by Mn(II) ($E_{a,VI} - E_{a,V} = -11 \pm 6$ kJ mol^{-1}).

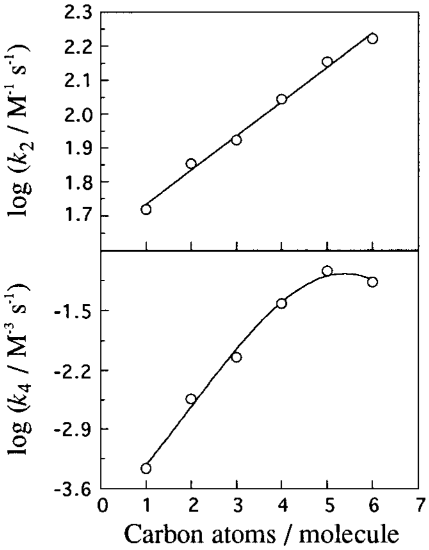


Figure 10. Dependence of the rate constants for the reduction of Cr(IV) (up) and Cr(VI) (down) on the number of carbon atoms per molecule of organic substrate at [$HClO_4$] = 0.366 M and 25.0 °C.

On the other hand, besides its inhibition effect on the initial rates, addition of Mn(II) to the solutions had a second effect on the reactions, since it resulted in an enhancement of their self-catalytic character. We suspect that this might be caused by a soluble form of colloidal manganese dioxide (known to have a self-catalytic effect on many permanganate reactions)[22] formed from the dismutation of Mn(III) and destroyed by a 2-electron reduction by the organic substrate:

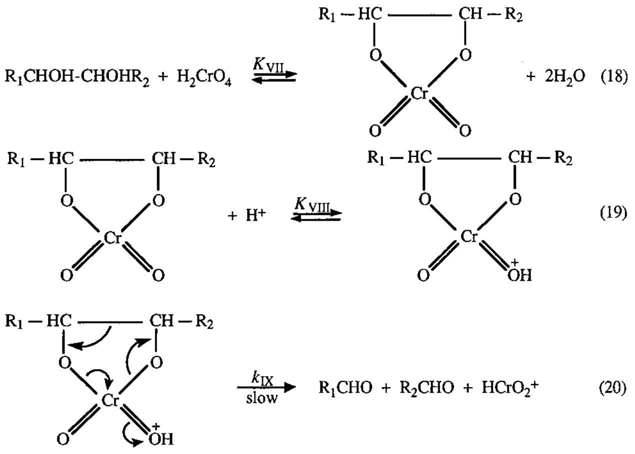
$$2\text{Mn(III)} \rightarrow \text{Mn(II)} + \text{Mn(IV)} \quad (16)$$

$$\text{Mn(IV)} + R_1R_2CHOH \rightarrow \text{Mn(II)} + R_1R_2C{=}O \quad (17)$$

Since the contribution of the intermediate Mn(IV) to the solution absorbance is caused by dispersion rather than by absorption of light, that contribution for a given colloidal particle size decreases as the wavelength of the radiation increases, and for a given wavelength it decreases as the colloidal particle size decreases.[23] Under the acidic conditions of our work, the reduction of Mn(IV) by the alcohol (eq 17) kept the size of the colloidal particles small enough, so that their contribution to the absorbance could be considered negligible. Actually, for the experiments done in the presence of Mn(II), well-behaved, perfectly uniform kinetic plots were obtained (Figure 1).

Oxidation by Cr(VI): Effect of the Alcohol Molecular Structure. The logarithm of the rate constant increased linearly with the number of carbon atoms of the alcohol for the oxidation by Cr(IV) (Figure 10, up), whereas for the oxidation by Cr(VI) a curve was obtained (Figure 10, down). A striking finding was that the reactions with Cr(VI) were much more sensitive to the alcohol molecular structure than their Cr(IV) counterparts. An increase of the number of carbon atoms from 1 to 6 resulted in a 150-fold increase in the rate of oxidation by Cr(VI) but only a 3-fold increase in the rate of oxidation by Cr(IV) (Table 3).

An explanation can be pointed out by considering the number of reaction pathways involved. In the case of the oxidations by Cr(VI), if we assume that only monoesters between the alcohol and chromic acid can be formed (as indicated in eq 7), the number of reaction pathways would equal the number of carbon atoms in the alcohol. However, if cyclic Cr(VI) diesters are also involved, the number of reaction pathways would increase. Their formation and destruction (illustrated for the case of 1,2-diesters) would be as follows:



The structure proposed for those 1,2-diesters keeps a close resemblance with that of the cyclic Mn(V) diesters proposed as intermediates for the oxidation of alkenes by permanganate both in aqueous media[24] and in organic solvents.[25]

The finding that the reactivity of the organic substrate toward Cr(VI) increased dramatically as its number of carbon atoms (and of alcohol groups, n) increased (Table 3) suggests that not only monoesters, but also 1,2-diesters, 1,3-diesters, ..., 1,n-diesters (up to the 1,6-diester in the case of D-mannitol) are formed between the alcohol and chromic acid. According to this hypothesis, the expression deduced from the proposed mechanism for the fourth-order rate constant corresponding to the oxidation of the alcohol of n carbon atoms by Cr(VI) in the absence of Mn(II) would be

$$k_{4,n} = 2\,K_I\Big(\sum_{i=1}^{n} K_{II,i}K_{III,i}k_{IV,i} + \sum_{i=1}^{n-1} K^{1,2}_{VII,i}K^{1,2}_{VIII,i}k^{1,2}_{IX,i} + \sum_{i=1}^{n-2} K^{1,3}_{VII,i}K^{1,3}_{VIII,i}k^{1,3}_{IX,i} + ... + K^{1,n}_{VII}K^{1,n}_{VIII}k^{1,n}_{IX}\Big) \quad (21)$$

where the first term between parentheses corresponds to the n reaction pathways involving monoesters, the second term to the $n - 1$ pathways involving 1,2-diesters, the third term to the $n - 2$ pathways involving 1,3-diesters, and successively so until the last term corresponding to the only pathway involving a 1,n-diester.

At 25.0 °C, the experimental rate constant for the oxidation of 1-propanol by Cr(VI) differs from that for the oxidation of 2-propanol in less than 10%.[5] Thus, for the same organic substrate and for Cr(VI) esters of the same type, the product of two equilibrium constants and one rate constant appearing in eq 21 right after each sum symbol is not expected to change very much from one alcohol group to another. Moreover, the equilibrium constant for the protonation of a chromium-bonded oxygen atom belonging to either a monoester (K_{III}) or a diester (K_{VIII}) intermediate is not expected to depend much on the particular reaction pathway considered. Hence, at least as a rough approximation, we can write

$$k_{4,n} \cong 2K_IK_{III}[nK_{II}k_{IV} + (n-1)K^{1,2}_{VII}k^{1,2}_{IX} + (n-2)K^{1,3}_{VII}k^{1,3}_{IX} + ... + K^{1,n}_{VII}k^{1,n}_{IX}] \quad (22)$$

One might attempt to go on with the approximations and to assume that the equilibrium constants for the formation of the esters (K_{II} or K_{VII}) and the rate constants for their redox destruction (k_{IV} or k_{IX}) are roughly independent of the reaction pathway and the alcohol considered. Since the total number of reaction pathways for the oxidation of the alcohol of n carbon atoms is $n(n + 1)/2$, the ratio of the experimental fourth-order rate constants for two consecutive alcohols in the homologous series should then be $k_{4,n}/k_{4,n-1} \cong (n + 1)/(n - 1)$, and the values expected for that ratio as n increases from 2 to 6 would successively be 3.0(6.7), 2.0(3.1), 1.7(4.2), 1.5(2.4), 1.4(0.7), the numbers between parentheses being the respective experimental values (Table 3). The latter are notably higher than the theoretical approximate values in all cases but for the last couple of alcohols.

Therefore, we must conclude that the products of equilibrium and rate constants appearing in eq 22 are actually strongly dependent on the particular reaction pathway and alcohol considered. Moreover, since in the homologous series studied each carbon atom supports an alcohol group and, due to its electron-withdrawing character, each OH group is expected to exert a negative effect on the reducing power of the other alcohol groups belonging to the same molecule of organic substrate (this point has been experimentally proved from Taft plots for the Cr(VI) oxidation of two different series of primary alcohols[20,21]), the rate constant for the rate-determining step of each reaction pathway (k_{IV} or k_{IX}) is expected to decrease as the value of n increases. Hence, the equilibrium constants corresponding to the formation of the ester intermediates (K_{II} or K_{VII}) must increase as n increases.

This hypothesis can be understood if we accept that the equilibrium constant for the formation of the monoesters is lower than those corresponding to the diesters, and that the latter increase as the separation of the two carbons supporting the chromate group increases too ($K_{II} < K^{1,2}_{VII} < K^{1,3}_{VII} < ... < K^{1,n}_{VII}$). The formation of a second C−O−Cr bond between the organic substrate and the oxidant would presumably render the structure of the ester more stable toward hydrolysis, because the electron withdrawing effect caused by the chromate moiety would be shared by two carbon atoms instead of being supported by just one. Moreover, the ester intermediates are expected to be more stable toward hydrolysis as the separation between the two carbon atoms bonded to the chromate group increases, due to the higher separation between the partial positive charges generated by that group on the two carbon atoms.

Summarizing, we explain the dramatic increase found for the reactivity of the alcohol toward Cr(VI) as the value of n increases by the combined result of two different effects: the increase of the number of reaction pathways involved in the mechanism and the possibility of formation of diester intermediates more stable toward hydrolysis. Thus, although it has been discussed if the Cr(VI) oxidation of diols takes place via acyclic chromate monoester or cyclic chromate diester intermediates,[26] our results with polyalcohols strongly suggest that both types of intermediate are involved in the mechanism.

The activation enthalpy decreased as n increased with two exceptions, corresponding to the oxidations of methanol ($n = 1$) and D-mannitol ($n = 6$) (Figure 11, up). The activation entropy increased linearly as n increased, but the value for methanol was much lower than expected from the linear plot (Figure 11, down). It is interesting to notice that the activation parameters obtained for the Cr(VI) oxidation of the monosaccharide D-ribose ($\Delta H^\circ_{\neq} = 40.1 \pm 0.6$ kJ mol^{-1} and $\Delta S^\circ_{\neq} = -155 \pm 2$ J K^{-1} mol^{-1})[10] are very close to the ones obtained

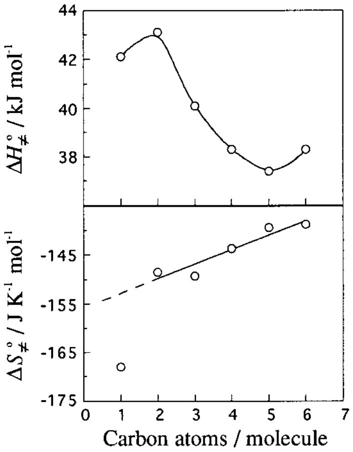


Figure 11. Dependence of the enthalpy (up) and entropy (down) of activation for the reduction of Cr(VI) on the number of carbon atoms per molecule of organic substrate.

in this work for the member of the homologous series with the same number of alcohol groups (DL-threitol, $\Delta H^{\circ}_{\neq} = 38.3 \pm 0.7$ kJ mol^{-1} and $\Delta S^{\circ}_{\neq} = -144 \pm 2$ J K^{-1} mol^{-1}), although both activation parameters are slightly less favorable for the reaction in the case of D-ribose. This difference might arise from the cyclic structure of the latter.

The enthalpy vs entropy of activation plot for the homologous series studied was not linear. This contrasts with the situation ordinarily found in chemical kinetics, since the enthalpy−entropy plots for most reaction families are linear, although some of those correlations might be provoked by the fact that the experimental errors in $\Delta H^{\circ}_{\neq}$ and $\Delta S^{\circ}_{\neq}$ for a particular reaction are not independent.[27,28] Moreover, for most reaction families there is a compensation effect between the activation parameters (either both $\Delta H^{\circ}_{\neq}$ and $\Delta S^{\circ}_{\neq}$ increase or both decrease).[29,30] However, for the Cr(VI) oxidation of the homologous series of alcohols studied in this work there seemed to be an anticompensation effect, since an increase in the value of n led to a decrease of $\Delta H^{\circ}_{\neq}$ in most cases, whereas $\Delta S^{\circ}_{\neq}$ increased (Figure 11). This can be understood on the basis of eq 22. Given that the factors n, $n - 1$, $n - 2$, ..., 1 corresponding to the number of reaction pathways involving each type of ester intermediate are independent of temperature, when applying the Eyring equation they appear lumped together in the value of $\Delta S^{\circ}_{\neq}$, thus counteracting the decrease of that magnitude usually associated to a decrease of $\Delta H^{\circ}_{\neq}$. Thus, the increase of $\Delta S^{\circ}_{\neq}$ as n increases is caused by the parallel increase in the number of reaction pathways involved, whereas the decrease of $\Delta H^{\circ}_{\neq}$ found in most cases when n increases is consistent with our hypothesis that an increase of the number of carbon atoms in the organic substrate opens the possibility for the reaction to follow more favorable reaction pathways.

The value of $\Delta H^{\circ}_{\neq}$ for the oxidation of D-mannitol ($n = 6$) seemed not to follow the general tendency of the family, since it was a little higher than that corresponding to adonitol ($n = 5$). Although such a small difference might be taken as the consequence of minor experimental errors, it is consistent with the finding that the rate constant for the oxidation of D-mannitol by Cr(VI) was considerably lower than that for adonitol (Table 3), although for all the other alcohols the reactivity increased as the value of n increased, and the difference in rate constants for $n = 5$ and 6 seems too high to be interpreted as an experimental error.

This may be explained by two alternative hypotheses. According to the first, due to its long hydrocarbon chain, D-mannitol might present a folded, stable conformation with an O−H hydrogen bond between the two terminal alcohol groups. Spectroscopic evidences of the existence of intramolecular hydrogen bonds between alcohol groups in the gauche conformation of ethylene glycol and in the DL-conformation of 2,3-butane diol have been reported.[26] If that hydrogen bond between the two terminal alcohol groups of D-mannitol does actually exist, the formation of the chromate ester intermediates involving one or both groups would require the cleavage of that hydrogen bond. This would be especially relevant in the case of the highly stable toward hydrolysis 1,6-diester, thus explaining the small increase in the value of $\Delta H^{\circ}_{\neq}$ for the oxidation of D-mannitol with respect to that for adonitol.

Alternatively, the increase in stability of the 1,6-diester intermediate for the oxidation of D-mannitol as compared with the 1,5-diester for adonitol might be of small quantity given that the separation between the two C−O−Cr bonds in the latter is already high enough. Hence, it might occur that the increase of the equilibrium constants for the formation of the diesters (K_{VII} in eq 22) when n increases from 5 to 6 was not high enough to compensate the decrease of the rate constants for the redox destruction of those intermediates (k_{IX}) caused by the accumulative electron-withdrawing effect of the OH group added.

The value of $\Delta H^{\circ}_{\neq}$ for the oxidation of methanol seemed also not to follow the general tendency of the family of reactions studied (Figure 11, up). More remarkably, the value of $\Delta S^{\circ}_{\neq}$ seemed exceptionally lower in comparison with those corresponding to the other five alcohols (Figure 11, down). This is in keeping with the reports that the rate constant for the Cr(VI) oxidation of methanol does not correlate well (too low) with those corresponding to two different series of primary monoalcohols,[20,21] and it might support the hypothesis that the redox destruction of the ester intermediate for the Cr(VI) oxidation of methanol follows a slightly modified reaction pathway with respect to the one applicable to the other alcohols.[21] Interestingly enough, the rate constant for the oxidation of methanol by Cr(IV) correlates very well with those corresponding to the reactions of other primary monoalcohols with the same oxidant.[5,31]

Oxidation by Cr(IV): Effect of the Alcohol Molecular Structure. Given that Cr(IV) is predominantly present in acidic aqueous solutions as the oxochromium(IV) ion, CrO^{2+},[5] that cannot form ester intermediates with the alcohols as chromic acid can, the second-order rate constant for the oxidation of the organic substrate of n carbon atoms by Cr(IV) in the absence of Mn(II) can be expressed simply as

$$k_{2,n} = \sum_{i=1}^{n} k_{\mathrm{V},i} \cong n k_{\mathrm{V}} \tag{23}$$

where the sum symbol is extended to the n reaction pathways corresponding to the Cr(IV) oxidation of each alcohol group, the rate constant for each pathway being $k_{\mathrm{V},i}$, and where it has been accepted as a rough approximation that, for a given organic substrate, the n values of $k_{\mathrm{V},i}$ are close enough to be replaced by an average value, k_{V}.

As inferred from Table 3, the second-order rate constant for the Cr(IV) oxidation of D-mannitol ($n = 6$) was only 3.2 times higher than that for methanol ($n = 1$). That ratio was lower than expected from the increase of reaction pathways. This finding can be explained considering that electron-withdrawing substituents such as the OH group exert un unfavorable effect on the reactivity of other alcohol groups present in the same reductant molecule, since they decrease the electron density on the carbon atom that must transfer a hydride ion (2-electron reduction) to Cr(IV) (eq 11).[2] Again, as in the case of Cr(VI), this point has also been experimentally confirmed for two different series of primary monoalcohols.[5,31] Hence, the value of k_V is expected to decrease as the number of OH groups in the reductant molecule grows.

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