Periodate–Glycol Reaction. 3. Pinacol, a Reevaluation

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The rate constants for the reaction of pinacol with periodic acid have been very accurately evaluated over the pH range 0.61–5.60 and at five temperatures 3.2–45 °C. At 3.2–25 °C the rate maximizes at pH 1.0–1.5; whereas at 35–45 °C the maximum rates are at pH 0.61. Both $\Delta H^{\ddagger}$ and $\Delta S^{\ddagger}$, the energy and entropy of activation, decrease markedly with decreasing pH at a given temperature and decrease modestly with increasing temperature at a given pH. It is concluded that periodic acid and pinacol undergo a direct bimolecular reaction but with the following considerations: (1) In addition to the major reaction of pinacol with periodic acid, there is a minor reaction with the periodate ion. (2) An additional equilibrium, $H_5IO_6 + H^+ \rightleftharpoons X^+$, deactivates periodic acid and causes a rate decrease at low pH, 25 °C or below. (3) The hydration equilibria of periodic acid and periodate ion appear to be largely responsible for the variations in $\Delta H^{\ddagger}$ with temperature and with pH. (4) Decreased hydration changes the concentration of the reactive form of the periodate ion or acid and thereby decreases $\Delta H^{\ddagger}$ with either decreased pH or increased temperature. (5) $\Delta S^{\ddagger}$ is determined from $\Delta H^{\ddagger}$ and is also greatly influenced by the extent of hydration. (6) The methyl groups of pinacol, as compared to ethanediol, retard the rate of reaction by steric effects but may enhance it by inductive effects.

Introduction

Periodate ion or periodic acid oxidizes vicinal glycols to form aldehydes or ketones and iodate ion or iodic acid. The oxidation is quantitative and proceeds at measurable rates. With ethanediol and certain other glycols there is clear kinetic evidence that a transitory complex is formed prior to final product formation. Although this complex has been generally assumed to be an intermediate,[1] this conclusion has been questioned.[2] Pinacol, 2,3-dimethyl-2,3-butanediol, exhibits no experimental evidence of such a complex.

There has been extensive early work on the periodate–glycol reaction, and this has been thoroughly reviewed.[3–5] The kinetics of the periodate–pinacol reaction were first studied by Price et al.,[6] and further work was done by Buist et al.[7] Rate maxima were shown to exist at about pH 1 and 8. No significant work has appeared since these publications.

There are some unusual and interesting facets to the periodate–pinacol reaction which have not been investigated, and several erroneous conclusions exist in the literature. The purpose of this paper is to present a more detailed and realistic view of this interesting reaction.

Experimental Section

The final concentration of KIO_4 used in all the rate studies herein reported was 0.002 000 M. The pinacol concentrations were either 0.004 826 or 0.004 736 M.

The reaction of pinacol with periodic acid was checked at pH 0.6 and 45 °C and was found to be quantitative in iodate formation with no reduction of iodate even after 2 h. These conditions were deemed the most likely for a possible continued reaction with iodic acid.

All kinetic data were determined by chemical procedures. Since the fastest reactions had rather short half-lives, it was necessary to employ the acid–iodide stop as previously described[2] and for the reasons indicated. In acid solutions periodates are reduced to iodine whereas glycols reduce periodate to iodate. The additional reduction necessitated a very large blank. This obvious disadvantage was offset by the development of a superaccurate and rapid analytical technique for iodine using silicone-coated autopipets and burets. The reproducibility was ±0.01%. The starting and sampling procedures as previously employed[2] were revised for greater efficiency and speed of handling. The data, reported herein, have been run three times at the lower three temperatures. All data were very consistent. The higher temperature studies were included for a more complete interpretation. The rate constants had an average deviation of ±0.2% except at 45 °C where the deviation was slightly greater.

Since the reaction was highly pH dependent, the control of pH was established with mild phthalate buffers for the higher pH and perchloric acid solutions at the lower pH. The same set of solutions was used at all temperatures. Measurements of pH were made using a PHM Radiometer pH meter at the five reaction temperatures except for the three lowest pH values. In order to obtain reasonably consistent results, several checks were made with each measurement, and then the overall process was repeated and in part repeated again. The pH measurements for each buffer solution were smoothed over the range of temperatures to give consistent results. Even with these precautions the measurement of pH provided the major uncertainty for the establishment of the accuracy of the data. This was especially true in the pH range 2–3.5, where the rate constants changed as much as 2% per 0.01 pH. The range of uncertainty in standardization of the pH meter was at least ±0.01 pH. Although measurements were read to 0.001 pH, the reliability was much less.

Since pH measurement at pH 0.6 gives uncertain results, the pH was estimated using the activity coefficient at 25 °C.[8] For pH 1.08 and 1.58, measurements were made only at 25 °C. Temperature adjustments were made by comparison to hydrochloric acid.

In order to adjust the data to the same pH and make it directly comparable at the five temperatures, the rate constants were adjusted to conform to a given pH as indicated using the factor $\Delta pH/\Delta k$ determined at each pH and temperature.

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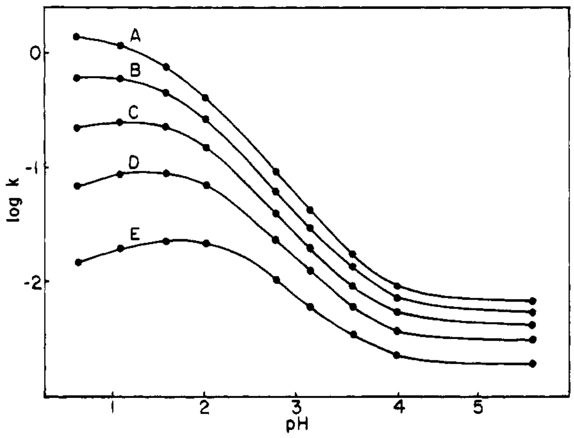

Figure 1. Plots of log *k* pH: A, 45.06 °C; B, 35.04 °C; C, 25.02 °C; D, 14.98 °C; E, 3.19 °C. The units for *k* are L/(mol s).

Pinacol was purified by crystallization from water as the hydrate. It was then dehydrated and further purified by zone refining. The freezing point was 43.1 °C. Potassium periodate, 99.99%, was from Aldrich Chemical Co. Deionized water, further distilled twice (once from $KMnO_4$), was used for all solutions.

Spectrophotometric measurements were made using a Hewlett Packard Diode Array Spectrophotometer HP8452 with a 1 mm quartz cell and 0.002 000 M KIO_4 at the indicated pH and temperatures.

Data and Discussion

Reaction rate constants have been determined in triplicate at nine pH values, each at five temperatures. The data, corrected as indicated in the Experimental Section, are graphed in Figure 1. As previously noted[7] and further verified for this work, no experimental evidence of complex formation between periodic acid and pinacol was observed.

The interpretation of these data must necessarily incorporate known equilibria

$$\begin{array}{ccccc} H_5IO_6 & \rightleftarrows & H_4IO_6^- + H^+ & \rightleftarrows & H_3IO_6^{2-} + 2H^+ \\ \upharpoonleft\!\downharpoonright & & \upharpoonleft\!\downharpoonright & & \\ HIO_4 & & IO_4^- & & \\ + 2H_2O & & + 2H_2O & & \end{array} \qquad (1)$$

Due to aberrant buffer effects in basic solutions, this paper is concerned only with acid solutions, pH 5.6 or less, so it is not necessary to take the $H_3IO_6^{2-}$ ion into consideration. The contribution of this ion is negligible at pH 5.6. However, as will be shown, yet another equilibrium exists at low pH.

The Question of Acid–Base Catalysis. Pinacol and ethanediol exhibit contrasting characteristics in their pH–rate data. Ethanediol[9] has shown a maximum over the pH range 3–6, and pinacol,[6,7] dual maxima at about pH 1–1.5 and pH 8 (0–25 °C).

With ethanediol the maximum rates are roughly in the pH range of the monovalent ion; accordingly it was concluded that the monovalent ion is the reactive species, and in order to support the principle of a single mechanism for this reaction, Duke and Bulgrin[10] followed by Buist et al.[7] have proposed that pinacol also reacts with the monovalent ion through the medium of acid–base catalysis. This proposal is reconsidered below.

For acid catalysis it may be assumed that P^-, the periodate ion, is the reactant with pinacol, b−x, and H^+ is the catalyst; then

$$dx/dt = k_p[H^+][P^-](b-x) \qquad (2)$$

Upon solving for P^- from (a−x) = [HP] + [P^-] and $K_i = [H^+][P^-]/[HP]$, where HP is periodic acid and *k* is a rate constant as designated,

$$dx/dt = k_pK_i(a-x)(b-x)[H^+]/(K_i + [H^+]) \qquad (3)$$

In contrast if it is assumed that HP is the specific reactant without catalysis, then

$$dx/dt = k_{HP}(a-x)(b-x)[H^+]/(K_i + [H^+]) \qquad (4)$$

The only difference between eqs 3 and 4 is in the labeling of the rate constant since $k_pK_i = k_{HP}$. Equation 4, however, is based on realistic considerations. A similar treatment in the basic region is also possible. Therefore the nebulous concept of acid–base catalysis may be discarded.

Evidence for the Dual Reactivity of Periodic Acid and Periodate Ion. Upon equilibrating eq 4 with the rate equation, $dx/dt = k_{obs}(a-x)(b-x)$; then

$$k_{obs} = k_{HP}[H^+]/(K_i + [H^+]) \qquad (5)$$

From the rate data used for Figure 1, $k_{obs} = 0.2501$ at pH 1.08 and 25 °C. With $K_i = 0.023$,[11] k_{obs} should be 3×10^{-5} at pH 5.6. This computed value is 0.1% of the experimental k_{obs}, which is 0.03083. Obviously dual reactions occur; thus, H_5IO_6 is the main reactant at low pH, and the monovalent ion is the predominant reactant near neutrality. There is gradation of the relative contributions of the two mechanisms as the pH varies from pH 0.6 to 5.6.

Protonic Equilibrium at Low pH. Note the peaks in the pH–rate constant graphs which occur at 25 °C and below. See Figure 1. Buist et al.[7] have attempted to explain the decreases in rate at high acidity by assuming that a transitory intermediate complex forms and becomes rate limited by high acidity. There is of course no evidence for this proposal, which appears to be based only on a predetermined assumption that the intermediate complex is a reality and has certain specified characteristics. The apparent success of their proposal is that three equilibria are involved, as needed to define the maximum in the rate curve.

Other explanations have been considered by the author. Under highly acid conditions pinacol rearranges to pinacolone, but even at pH 0.6 the acidity appears to be too low to effect this reaction. Such was demonstrated by allowing a solution of pinacol to stand at pH 0.6 for about 1 h before reacting with periodic acid. No change in rate was detected. Upon adding sodium perchlorate at pH 1.0 to provide an ionic strength equal to that of perchloric acid at pH 0.6, there was a small increase in rate. Thus, it was further determined that a salt effect was not causing the rate decrease.

The possibility of an equilibrium between H_5IO_6 and H^+ was then investigated. Spectrophotometric observations were made in the range 190–250 nm with 0.002 M KIO_4 at the indicated pH and temperature values. Dual peaks show an isosbestic point at 204.5 nm which is essentially maintained over the entire temperature range. One peak maximizes at 222–224 nm, and the other peak maximizes below 190 nm. The data have been plotted, absorbance versus pH, in Figures 2 and 3.

The IO_4^- ion strongly absorbs at 222 nm. A plot of the absorption by 0.002 M KIO_4 at the varying pH and temperature values is seen in Figure 2. There is a decrease in absorption with both decreasing pH and decreasing temperature. The

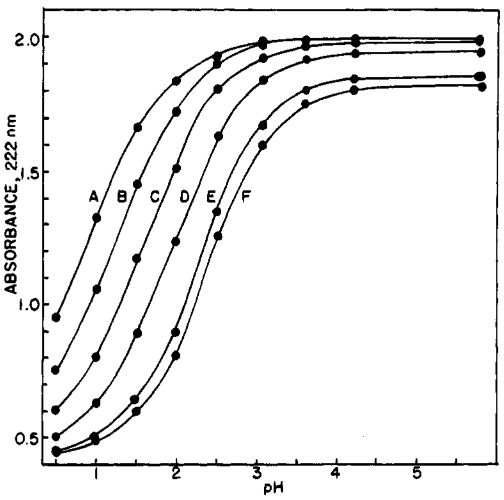

Figure 2. Plots of absorbance vs pH at 222 nm: A, 45 °C; B, 35 °C; C, 25 °C; E, 4 °C; F, 0.5 °C.

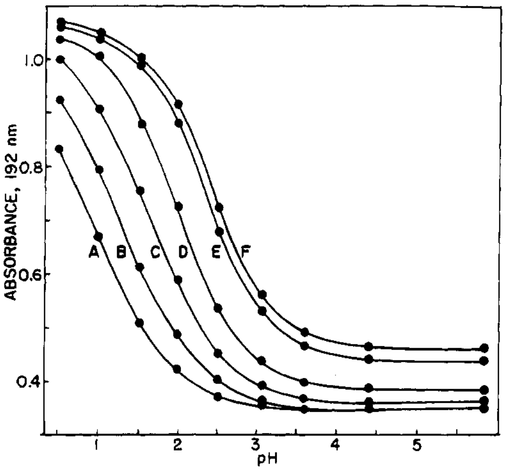

Figure 3. of absorbance vs pH at 192 nm: A, 45 °C; B, 35 °C; C, 25 °C; D, 15 °C; E, 4 °C; F, 0.5 °C.

former is due to increased conversion of IO_4^- to H_5IO_6, and the latter, to increased hydration of IO_4^- to $H_4IO_6^-$.

A contrasting phenomenon occurs with 0.002 M KIO_4 at 192 nm, where absorption by the IO_4^- is very small. As seen in Figure 3 absorption at this wavelength increases both with decreasing pH and with decreasing temperature. Obviously a product is formed which is an adduct of H^+ and H_5IO_6. (There also appears to be an additional absorbance.) This product has not been identified and could be $H_6IO_6^+$ or a dimeric or polymeric species. Accordingly it will be designated as X^+.

$$H^+ + H_5IO_6 \rightleftarrows X^+ \text{ and } K_X = [H_5IO_6][H^+]/[X^+] \quad (6)$$

Mishra and Symons[12] also made spectral observations and described X^+ as $H_6IO_6^+$. X^+ is obviously of low reactivity, which explains the decrease in the rate constant values as seen in Figure 1.

Evaluations of the Constants. In order to estimate the contributions of various equilibria at each of the five temperatures, the following equation was derived without approximation.

$$k_{\text{calc}} = \frac{k_1 K_i K_X + k_2 K_X [H^+] + k_X [H^+]^2}{K_i K_X + K_X [H^+] + [H^+]^2} \quad (7)$$

where k_1, k_2, and k_X are the maximum rate constants for the respective reactions of pinacol with P^-, HP, and X^+; K_i is the first ionization constant, and K_X is from eq 6.

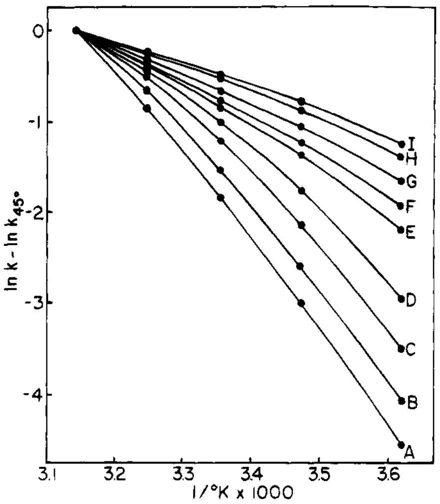

Figure 4. k minus ln k at 45 °C values plotted for each pH to show the consistency of the data. pH values: A, 0.61; B, 1.08; C, 1.58; D, 2.02; E, 2.78; F, 3.15; G, 3.62; H, 4.10; I, 5.60.

Values for k_X, k_2, and K_X were determined using k_1 as observed at pH 5.6, and K_i was from the literature[11] or estimated. A least squares computer program was then applied, giving estimates of k_2 (L/(mol s), K_X, and k_X at the following temperatures: (3.29 °C) 0.027, 0.26, 0; (14.98 °C) 0.12, 0.35, 0; (25.02 °C) 0.33, 0.35, 0; (35.04 °C) 0.77, 1.2, 0; (45.06 °C) 1.5, very large, 0. Comparable experimental maximum values for k_2 are 0.022 57 at 3.29 °C and pH 1.58, 0.0884 at 14.98 °C and pH 1.58, 0.2504 at 25.02 °C and pH 1.08, 0.602 at 35.04 °C and pH 0.61, and 1.386 at 45.06 °C and pH 0.61.

The complete nonreactivity of X^+ is clearly indicated above. The temperature dependence of X^+ is shown by the very large value of K_X at 45 °C, indicating the virtual disappearance of X^+. This is further supported by the decreasing values of (k_2 experimental − k_2 computed) with temperature, which become nearly zero at 45 °C as seen above.

Activation Energies and Entropies. To demonstrate consistency, a plot of the rate data, as ln k vs $1/T$, has been made. The ln k value at 45 °C has been substracted from each set of data at a given pH. This has been done to avoid the confusing set of crossed curves which results from a simple ln k vs $1/T$ plot. It is important to emphasize the excellent consistency of these data in order to give validity to any interpretations to be made. See Figure 4.

A somewhat unusual phenomenon is exhibited by this reaction in that the low activation energy and low rate at pH 5.6−3.6 change to both a greater activation energy and greater rate at pH 1.5−0.6. Larger than equivalent changes in entropies of activation offset the activation energy increases.

The energies $\Delta H^\ddagger$ and entropies $\Delta S^\ddagger$ of activation have been calculated and are plotted in Figures 5 and 6. Obviously these data are inconstant and seemingly inconsistent where constancy and consistency are expected.

The extent of hydration of a reactant changes the rate of a reaction by altering the effective concentration of that reactant. When there is a change in the extent of hydration with temperature, pH, or any other factor, then the calculated activation energy will also vary with temperature. This has been previously discussed.[2] Thus the activation energies and entropies graphed in Figures 5 and 6 represent composite numbers consisting of the fundamental activation energy plus extraneous additions. A qualitative evaluation follows listing the most probable causes of these inconsistencies.

1. The extent of hydration of the periodate ion and periodic acid increases with decreased temperature. Correspondingly,

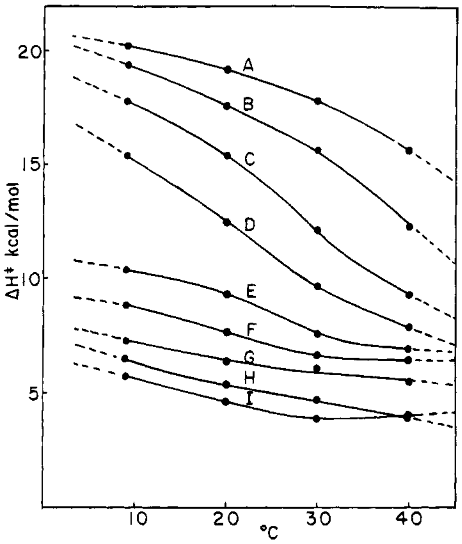

Figure 5. Plots of computed $\Delta H^\ddagger$ values at each pH over a range of five temperatures. The dotted lines are extrapolated. pH values: A, 0.61; B, 1.08; C, 1.58; D, 2.02; E, 2.78; F, 3.15; G, 3.62; H, 4.10; I, 5.60.

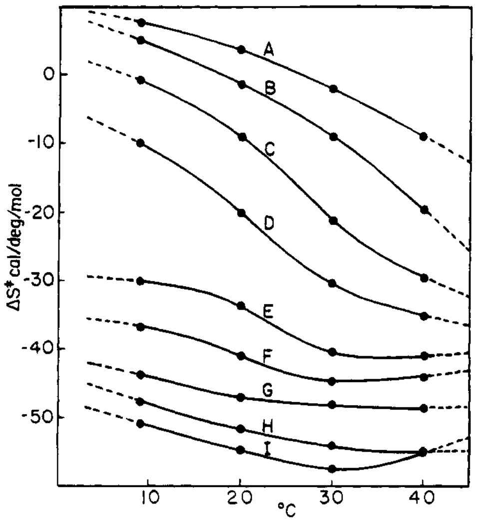

Figure 6. Plots of the computed $\Delta S^\ddagger$ values at each pH over the range of five temperatures. pH values: A, 0.61; B, 1.08; C, 1.58; D, 2.02; E, 2.78; F, 3.15; G, 3.62; H, 4.10; I, 5.60. The lines for temperatures 9.1–3.2 °C and 40–45 °C are extrapolated.

both $\Delta H^\ddagger$ and $\Delta S^\ddagger$ increase with increasing temperature. A similar phenemenon was previously described for the monovalent periodate ion in conjunction with the periodate–ethanediol reaction.[2]

2. For periodic acid the extent of hydration is also related to the extent of ionization. Periodic acid, H_5IO_6, is largely hydrated, and the monovalent ion, IO_4^-, is largely unhydrated, particularly at higher temperatures. Therefore with decreasing pH there is increasing conversion of the ion IO_4^- to the acid, and both $\Delta S^\ddagger$ and $\Delta H^\ddagger$ increase with a decrease in pH.

3. The formation of X^+ is enhanced by decreased pH and decreased temperature. Since X^+ is inactive the rate of reaction is lowered with decreased pH. In turn the apparent values of $\Delta H^\ddagger$ and $\Delta S^\ddagger$ are altered. With an increase in temperature the concentration of X^+ converts back to H_5IO_6. This has the effect of increasing the rate in excess of the normal increase due to temperature; i.e., the rate change is greater than if X^+ had not formed at the lower temperature. These extrinsic changes in $\Delta H^\ddagger$ and $\Delta S^\ddagger$ are verified in Figure 5, since the highest values are at pH 0.61.

4. In basic solutions a marked buffer (or anionic) effect has been noted. This may possibly carry over to a limited extent into low acid solutions containing the phthalate ion.

5. The method of calculation of E_{act} and $\Delta H^\ddagger$ from $\Delta \ln k$ and $1/T_2 - 1/T_1$ can give inconsistencies if the curvatures of the rate constants with temperature are not consistent. In both Figures 5 and 6 a median temperature was chosen as the point at which each $\Delta H^\ddagger$ or $\Delta S^\ddagger$ is identified, but this serves only as an estimate of the slope at that point.

6. The corresponding large changes in $\Delta S^\ddagger$ undoubtedly reflect the changes in $\Delta H^\ddagger$ since $T\Delta S^\ddagger = \Delta F^\ddagger - \Delta H^\ddagger$; generally as $\Delta H^\ddagger$ increases with decreasing pH and temperature, so does $\Delta S^\ddagger$. The reaction at pH 5.6 has a very low $\Delta S^\ddagger$, −50 to −57 eu, but with the decrease in pH it increases to +8 to −9 eu. Thus the lower values of $\Delta S^\ddagger$ at pH 5.6 compared to those at pH 0.6 also correlate with the change from a largely unhydrated reactant at pH 5.6 to a hydrated reactant at pH 0.6.

7. Partial dehydration of H_5IO_6 occurs at higher temperatures, as implied from both spectral and rate data. Absorption of IO_4^- has been followed at 222 mm in the pH range 5.6. See Figure 2. It has also been observed that there is an increase in absorption with increasing temperature at 222 nm in periodic acid solutions at pH 0.6, which suggests that partial dehydration of H_5IO_6 occurs with increased temperature. This correlates with the rate data, since a decrease in activation energy is noted with increasing temperature.

8. The heat of dehydration of $H_5IO_6 \rightleftharpoons 2H_2O + IO_4^- + H^+$ has been estimated as $\Delta H = 10.9$ kcal/mol,[11] which is the same as that for $H_4IO_6^- \rightleftharpoons 2H_2O + IO_4^-$. In view of the greater stability of H_5IO_6 compared to $H_4IO_6^-$, the ΔH value for H_5IO_6 is low or else there are differences in the ΔS of dehydration for H_5IO_6 and $H_4IO_6^-$.

Mechanistic Proposals

The primary mechanism appears to be a simple bimolecular reaction whereby periodic acid reacts directly with pinacol. The products are formed instantaneously after the initial contact. The mechanism which then occurs remains to to be determined. The complexities are brought about by various factors, as previously discussed.

The products of the oxidation of pinacol by periodic acid, HIO_3 and acetone, are unhydrated whereas the apparent reactant H_5IO_6 is strongly hydrated. Whether this loss of water occurs before, during, or after formation of the activated complex provides a basis for speculation which will be further considered in a future publication.

As previously described, both H_5IO_6 and $H_4IO_6^-$ appear to react with pinacol, but $H_4IO_6^-$ is much less reactive than H_5IO_6. This seems to imply dual mechanisms, but the nature of the differences in the mechanisms is not defined. Scott and Tigwell[13] have demonstrated that using nonhydroxyl solvents such as dioxan greatly increases the rate of reaction of pinacol with periodic acid as compared with water. Although further study of the significance of these observations is needed, support for the duality of mechanisms is implied, since it seems unlikely that the periodate ion would show increased reactivity in a nonaqueous solvent.

Steric and Inductive Effects. The previous discussion has dealt exclusively with periodic acid. The other reactant, pinacol, must also be evaluated. The four methyl groups of pinacol, which differentiate it from ethanediol, are an important consideration in the reaction. The methyl groups act as barriers to the rotation about the central C–C axis and also as a steric effect to hinder contact between the reactants. In contrast the methyl groups can presumably increase the polarity of the

C–OH bonds and weaken the central C–C bond where the scission occurs. Upon comparing the present data with earlier results,[2,9] ethanediol at pH 5.6 has a rate constant about 600 times greater than that of pinacol, but at pH 0.6 the rate constants are only about 2 times greater than that of pinacol. Dual effects, one to decrease the rate and another to enhance it, are a possibility.

Steric hindrance by the methyl groups is obvious from the rate differences as noted above. There are two possible approaches of periodic acid to the pinacol molecule. The first is either a rearward or side approach to the glycol carbons. In view of the blockage of approach to the central carbons by the methyl groups, this appears to be an unlikely possibility. The second is in keeping with the traditional proposal[1] whereby periodic acid reacts with the hydroxyls to form the products. The steric effect by the methyl groups to retard the reaction could be due to a hindered rotation of the central C–C bond, making more difficult the proper orientation of the hydroxyls for interaction with periodic acid. Spatial hindrance of this interaction by the methyl groups may also be a factor.

Enhancement of the rate by the methyl groups is conceivable, since there should be an inductive effect to increase polarity of the C–O bonds and also to weaken the central C–C bond. To what extent this may increase the rate of reaction is not known.

More complete and accurate determinations of the thermodynamic constants related to periodic acid over the range of temperatures of this study would permit a more quantitative determination of the contribution of periodic acid and the related hydration phenomena. The other forces, such as steric effects, which also contribute to this reaction could then be more accurately evaluated.

Conclusion

This work provides an argument against popular attempts to simplify and classify reaction mechanisms. As is shown in this paper, the mechanism of any reaction is subject to minor forces which can greatly alter the course of a reaction. Thus it is not possible to make general correlations without detailed considerations of qualifying conditions.

Acknowledgment. The author is greatly indebted to Dr. John Reed, who developed the computer programs, and to Virginia Carrick and Christina Sauer, who did preliminary studies on this reaction.

Supplementary Material Available: The original rate data, a detailed description of the revised kinetic apparatus, analytical procedures, buffers used, the derivation of eq 6, and figures showing the automatic precision pipet and the rate starting and stopping apparatus (9 pages). Ordering information is given on any current masthead page.

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