

OXIDATION OF ORGANIC COMPOUNDS

WITH PENTAVALENT VANADIUM

A Thesis submitted for the Degree of

Doctor of Philosophy

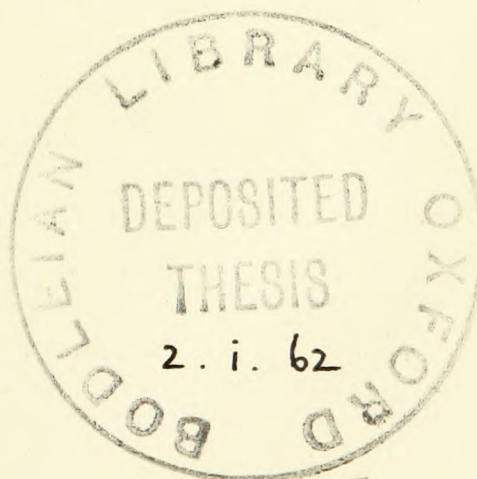
in the University of Oxford.

by

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CHAPTER I

INTRODUCTION

Solutions of compounds of quinquivalent vanadium (V^V) in acid have been known to possess oxidising properties ever since Selstrom attempted to isolate metallic vanadium in 1830. Berzelius¹ in his account of this work describes the reduction by alcohol of a dark red solution of impure vanadium oxides in concentrated sulphuric acid to a blue solution. The use of V^V as an oxidising agent has however been very limited and confined largely to analytical purposes. A brief account of the history of the various uses of V^V as an oxidising agent has been given by J.S.Littler². The only kinetic study of V^V oxidations made until recently was that of Bobtelsky and Glasner³ who oxidised arsenious oxide, oxalic acid, formaldehyde, and ethanol in sulphuric acid solutions and also observed the effect of various inorganic anions and cations on the rate of reaction. Their work will be described more fully later.

Littler and Waters⁴ have now surveyed the oxidation of various organic compounds by V^V and have made kinetic measurements on a number of simple alcohols, ketones and glycols. Polymerisation of acrylonitrile and reduction of mercuric chloride were induced by most oxidations,

effects which were taken to demonstrate the presence of free radicals in the reaction and hence of one-electron transfer steps in the oxidation processes. These are necessary if V^V is only reduced to V^{IV} ; no evidence for V^{III} in the oxidation mixtures has yet been found. For the oxidation of pinacol in sulphuric acid solutions they deduced the equation

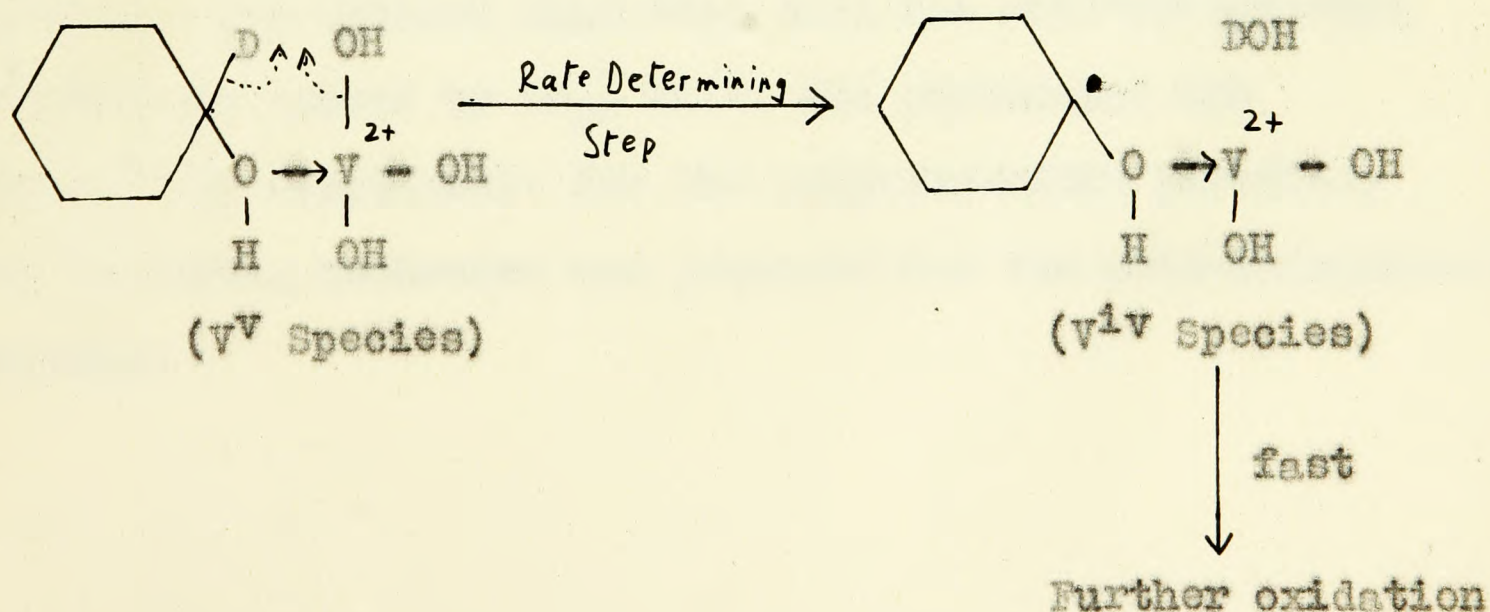
$$\frac{-d[V^V]}{dt} = k_0 [Pinacol] [V^V] (1 + f[H^+])$$

where $f[H^+]$ is an undetermined function of acidity; the observed rate increased sharply as the acidity was increased above 1.0N. Pinacol was cleaved quantitatively to acetone.

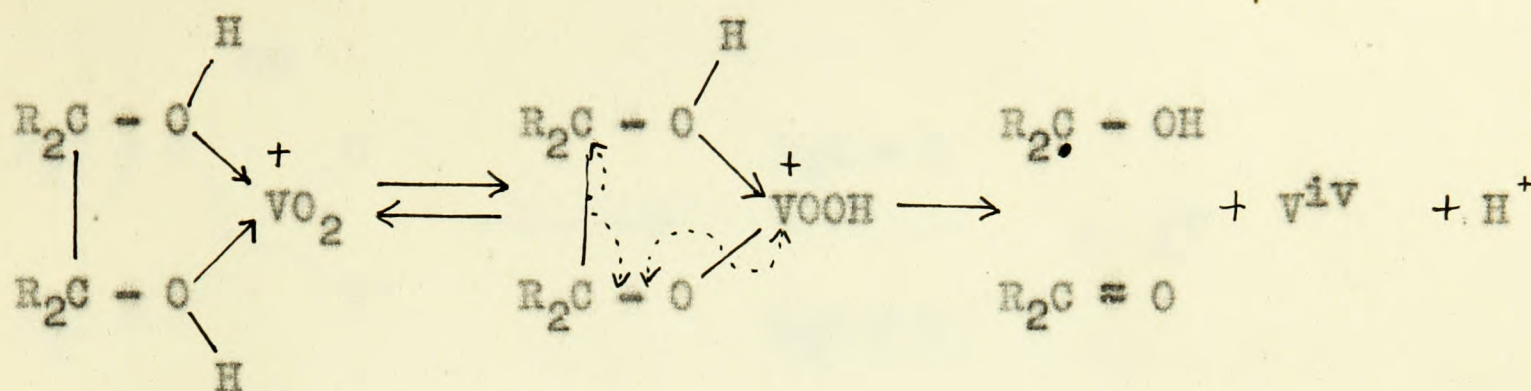
The oxidations of cyclohexanone and of cyclopentanone were found to be first order reactions with respect to both V^V and ketone and an acid catalysis effect was again found. The rate of oxidation in perchloric acid solutions of constant ionic strength was directly proportional to acid concentration $[H^+]$ but in sulphuric acid solution the rate increased more rapidly with acid concentration. This was attributed to the powerful oxidising properties of sulphate complexes of V^V which are formed in sulphuric acid of high concentrations. Oxygen was found to retard the oxidation of cyclopentanone,

but although it was absorbed it had no effect on the rate of the cyclohexanone oxidation. The product formed from cyclohexanone was adipic acid and from cyclopentanone the products were probably an equimolecular mixture of succinic and formic acids.

Cyclohexanol was also oxidised to adipic acid, but more slowly than cyclohexanone, so that the presence or absence of cyclohexanone as an intermediate could not be shown. The kinetics of the reaction again were of first order in substrate, V^V and acid concentration. 1-Deuterio-cyclohexanol was found to be attacked more slowly than the protio compound indicating C - D bond breaking in the rate-determining step. It was also shown by ultra-violet absorption spectra that the alcohol formed a complex with the vanadium species. The following mechanism was proposed



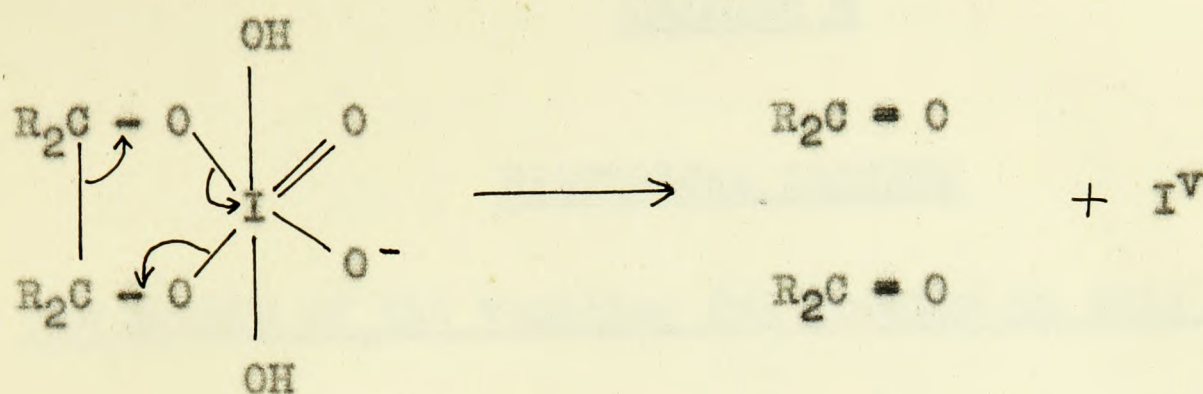
Littler, Mallet and Waters have oxidised some α -glycols and found a first order rate dependence on substrate and V^V . The acidity dependence was less consistent however. Glycols containing only primary or secondary hydroxyl groups were oxidised like cyclohexanol with a first order dependence on acid. However glycols with a tertiary hydroxyl group, such as cis and trans 1.2-dimethyl cyclohexane 1.2-diol and 2-methylbutane 2.3-diol, like pinacol, were oxidised with both an acid independent reaction and an acid catalysed reaction, the latter increasing more rapidly with acid concentration than would be the case for a first order reaction with respect to hydrogen ion concentration. Moreover, pinacol and 2-methylbutane 2.3-diol were oxidised much faster than the other glycols, pinacol reacting particularly rapidly. Cyclic cis glycols were oxidised faster than the corresponding trans glycols. The authors suggested that the acid-independent reaction is caused by VO_2^+ , while the protonated ion $V(OH)_3^{2+}$ is responsible for the acid-dependent reaction. The following mechanism was proposed for the acid-independent reaction



I

The cyclic complex I imposes steric limitations on the type of glycol which can be most readily attacked. When a radical $\text{R}_2\text{C}^\bullet - \text{OH}$, which is more stable than $\text{RCH}^\bullet - \text{OH}$, can be formed, then C - C fission occurs in preference to a reaction of the type which has been proposed for the oxidation of cyclohexanol which requires $\text{V}(\text{OH})_3^{2+}$ as the oxidising species. For this reason tertiary glycols behave differently to primary and secondary glycols with respect to V^{V} .

In contrast, the oxidation of glycols by a two - electron oxidising agent such as periodate can proceed through a cyclic intermediate which depends only on steric requirements and not on the presence of a tertiary hydroxyl group.



II

The complex II must have a planar ring, whereas that proposed for the V^V oxidation I may be puckered.

CHAPTER 2

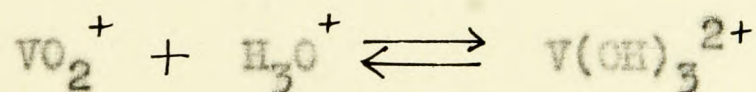
HISTORICAL SECTION

The Nature of the Vanadium (v) Species in Acid Solutions

Vanadium (v) in aqueous solutions is amphoteric, the isoelectric point being at an acidity of about $\text{pH} = 2$. Its oxidation potential⁶ is $+0.9$ volts at $\text{pH} = 1$, increasing to over 1.3 volts in 10 N acid and decreasing to -0.4 volts at $\text{pH} = 10.5$. Thus V^{V} only possesses oxidising properties in acid solutions.

The principal V^{V} species present in dilute solutions of mineral acid was at one time thought⁷ to be VO^{3+} , but it is now generally agreed^{6,8,9} to be VO_2^+ . The degree of hydration of this cation is not certain, but entropy measurements¹⁰ suggest that the ion is probably VO_2^+ aq. rather than $\text{V}(\text{OH})_4^+$ aq.. In order to account for the observed acidity dependence of the oxidation rate of various organic compounds, Littler and Waters⁴ have proposed that in more concentrated acid solutions, the more strongly oxidising cation $\text{V}(\text{OH})_3^{2+}$ is formed to a small extent. The visible and ultra violet spectra of aqueous solutions of V^{V} do not change appreciably with increasing concentrations of perchloric acid (an acid which has no complexing power towards V^{V}). The spectrum might

have been expected to change if a different cation is formed to some extent at high acidity, but the absence of such a change does not preclude the existence of $V(OH)_3^{2+}$. In a wide variety of oxidations at constant ionic strength, the rate of reaction of V^V has been found to be directly proportional either to the first power of the hydrogen ion concentration, or to the acidity function h_0 (see Chapter 4). If it were the organic substrate that was being protonated rather than the V^V species, some variation of the acidity dependence would have been expected. Also the observed protonation occurs at acidities lower than those at which alcohols and ketones etc. are normally protonated to any extent. The solvent isotope effects in several V^V oxidations have given a ratio of rate constants k_{D_2O}/k_{H_2O} which is in accordance with the occurrence of a pre-equilibrium protonation of a reacting species¹¹, and is consistent with an ionic equilibrium



(The theoretical basis of solvent isotope effects is discussed later.)

West and Skoog¹² have postulated the species VO^{3+} in strong acid solution as they found a second order dependence on acidity in the V^V oxidation of glycerol in sulphuric acid solutions. Since they do not mention whether their

ionic strength was kept constant, it is possible that the reaction may have followed h_0 , which roughly parallels $[H^+]^2$ in the region of low acidity, and that a positive salt effect accelerated the reaction at higher acidities. In any case V^V oxidations in sulphuric acid are complicated by the presence of V^V - sulphate complexes, which give strong acid solutions a deep red colour and possess very strong oxidising properties. Various other inorganic anions are also capable of increasing the intensity of the pale yellow colour of VO_2^+ , but the perchlorate anion apparently has no tendency to complex with V^V , and for this reason reaction kinetics obtained in solutions containing only perchlorate anions are more readily interpreted than results obtained in, say, sulphuric acid solutions.

Methods of Determining Reaction Mechanisms

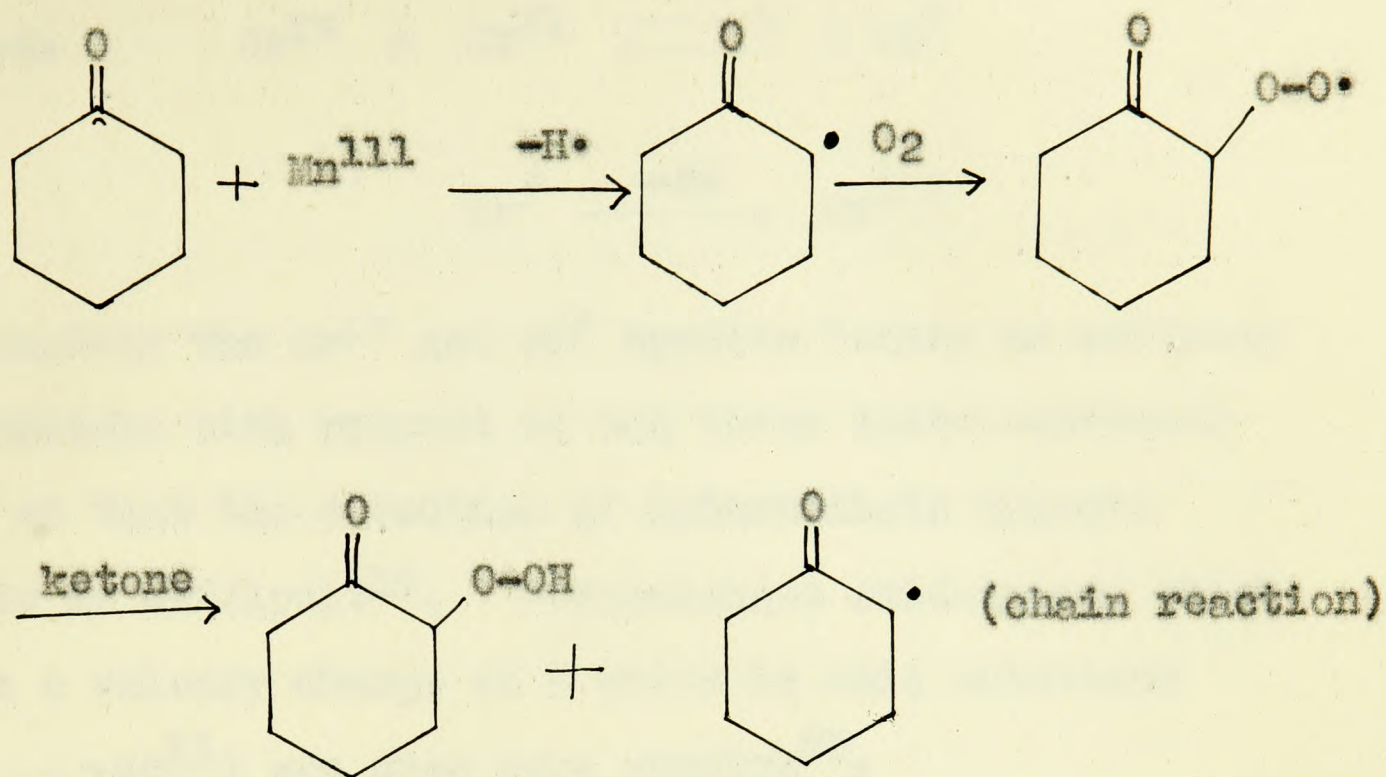
A great deal of information about the mechanism of a reaction can be obtained from an analysis of the products and from a detailed investigation of the kinetics of the reaction under varying conditions of concentration of reactants, temperature, pressure etc.. Investigations of the catalytic effects, whether positive or negative, of various added materials, such as mineral acid and selected inorganic ions, can often lead to conclusions about the rate determining step in a reaction. The type of salt effect observed may give some indication of the nature of an intermediate complex or transition state, and particularly of the charge on such a species¹³. There are a number of other weapons which can now be employed in attempting to elucidate reaction mechanisms and some of these which are relevant to oxidation reactions are described here.

The Investigation of Reactions Involving Free Radicals

Various workers¹⁴ have attempted to divide oxidising agents into two types - the so-called one-electron and two-electron oxidants, according to the number of electrons

transferred between oxidant and reductant in the oxidation step. One-electron transfers from organic molecules often result in the formation of organic free radicals, but occasionally the radicals never get quite free from the reacting complex before a subsequent reaction occurs.

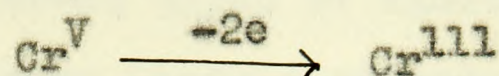
Free radicals can be detected in oxidation mixtures by various techniques. The polymerisation of acrylonitrile is normally initiated in situ by mixtures containing free radicals¹⁵; the polymer is insoluble in water and precipitates out of solution. The nature of the initiating radical, which is incorporated in the end of each polymeric molecule, can be investigated by infra red spectroscopy. Mercuric salts are reduced to the mercurous state by many organic radicals¹⁶. Induced reductions of other materials are also observed, e.g. Iron (III) $\longrightarrow$ Iron (II)¹⁷. Not all free radicals are reducing agents. For instance the radical $\text{HOOC}\cdot\dot{\text{C}}\text{H}\cdot\text{COOH}$, derived from malonic acid has two powerful electron withdrawing groups and is an oxidising agent rather than a reducing agent¹⁷. Radicals with the odd electron on an oxygen atom are also oxidising agents and have no effect on mercuric salts. Oxygen can often be absorbed in reactions involving free radicals, e.g. in the manganese (III) oxidation of cyclohexanone¹⁸



In general oxidations by vanadium (v) show these three characteristic effects and, since no vanadium (III) has yet been detected during any oxidations, reactions may be assumed to proceed through one-electron transfer steps:-

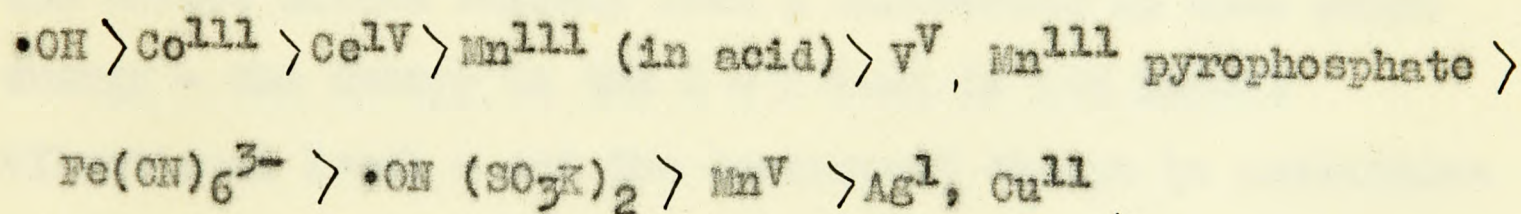


In the case of chromic acid oxidation when the valency state changes by more than 2 units a more complicated series of steps must occur; these may be all one-electron, a mixture of one- and two- electron steps or all two-electron steps as far as the organic reactants are concerned for disproportionation of the intermediate valency states of chromium may occur.

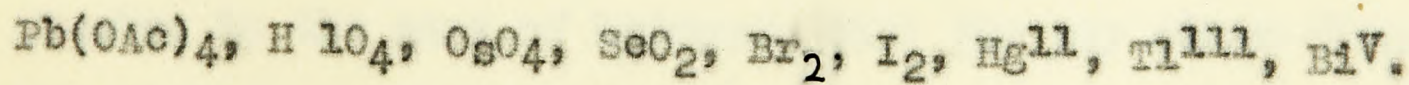


Unfortunately the Cr^{IV} and Cr^{V} species behave as reducing free radicals with respect to the three tests mentioned above, so that the detection of intermediate organic radicals is difficult¹⁹. Permanganate oxidations, which involve a valency change of 5 units in acid solutions ($\text{Mn}^{\text{VII}} \longrightarrow \text{Mn}^{\text{II}}$) are even more complex²⁰.

In general the following oxidants appear to act as one-electron transfer reagents and are given in order of decreasing reactivity (which corresponds closely to the order of their redox potentials):-



The following are two-electron transfer reagents:-



These are not given in any particular order since the oxidative power of each reagent varies according to the substrate used, owing to the diversity of the different mechanisms involved.

Isotope Effects

There are three different types of kinetic hydrogen isotope effects: - A. primary isotope effects, in which a bond to a hydrogen or deuterium atom is broken; B. secondary isotope effects in which the isotopic atoms are not directly involved in the breaking or formation of bonds but may act inductively; and C. solvent isotope effects which include all the effects of replacing H₂O by D₂O as the solvent, and which may include examples of primary and secondary isotope effects.

A. Primary Isotope Effects.

These always give values for the ratio of the rate constants k_H/k_D^* > 1 and usually between 1.5 and 7.0. The effect arises largely from a difference in zero point energy - the energy in the C - H bond in its lowest vibrational level - and the consequent change in activation energy to produce bond rupture. This difference is related to the ratio of the masses of the two isotopes and is greater in the case of hydrogen than for any other element. The variation in the observed values of k_H/k_D is due to the nature of the transition state which is formed with the

* Subscript refers to isotopic form.

reacting species and which will result in the C - H (C - D) bond being stretched to a greater or smaller amount in different cases prior to the bond actually breaking. The minimum value of a primary deuterium isotope effect is $\sqrt{2}$ which is calculated from the different mass of the H and D atoms only²¹. An allowance must also be made for non-classical penetration of the energy barrier ("tunnelling"), which is more likely to occur with hydrogen than with deuterium and which enhances the k_H/k_D ratio at low temperatures²².

For C - H and C - D bonds the maximum values of k_H/k_D without tunnelling have been calculated as follows²³:-

Temp. °C	0	25	100	200	300	500
k_H/k_D	8.3	6.9	4.7	3.4	2.7	2.1

In mechanistic studies the importance of rate measurements of primary isotope effects is that they show whether a particular X - H bond is, or is not, broken in the rate determining step of a reaction.

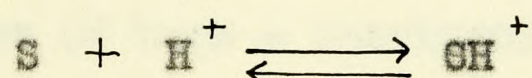
B. Secondary Isotope Effects

These are usually small and k_H/k_D may be positive or negative. The C - D bond may influence reaction rates

(i) by its greater inductive electron-releasing ability²⁴, (ii) by its effect on the overall vibrational energy and thus on the entropy of the molecule²⁵, (iii) by its reduced ability to stabilise the transition state by hyperconjugation²⁶ (iv) but its hybridisation being changed in the transition state²⁷, and (v) by its reduced power of participating in non-classical carbonium ion intermediates²⁸.

C. Solvent Isotope Effects

The most important effect is that on acid - base catalysis. Many acid catalysed reactions proceed more rapidly in D₂O than in H₂O. This can be explained²⁹ if these reactions are those which involve a rapid pre-equilibrium between the substrate and hydrogen ion:-



followed by a slow decomposition of SH⁺ to products. The latter step will involve a (small) secondary isotope effect and possibly a primary isotope effect. Neglecting the secondary isotope effect the rate will depend on the concentration of SH⁺ or SD⁺. The acid dissociation constant of SH⁺ in H₂O is greater than that of SD⁺ in D₂O by a factor changing from 2.74 to 4.16 for acids with pK_a

changing from 3 to 11. Values of pK_a for the protonated forms of substrates which are involved in the type of reaction under consideration (e.g. acetone), are much less than 3 and are generally negative. The factor k_{D_2O}/k_{H_2O} would consequently be expected to be rather less than 2.74. Experimental values are usually about 2, which substantiate the above theory of the origin of the solvent isotope effect.

In some cases larger values of k_{D_2O}/k_{H_2O} are observed owing to double protonation of the substrate.

If the species SH^+ is a weak acid most of the substrate is present in the protonated form and the relative rates of decomposition of SH^+ and SD^+ control the observed isotope effect which may then be either positive or negative.

In the case of base - catalysed reactions, catalysed by OH^- (OD^-), the rate of reaction in D_2O is faster because the basic strength of OD^- is greater than that of OH^- .

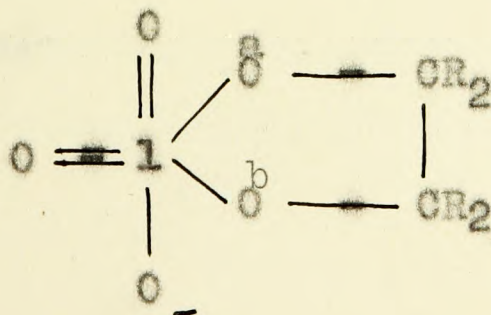
Under carefully controlled conditions, solvent isotope effects can be just as useful in elucidating reaction mechanisms as primary isotope effects.

Other Uses of Isotopes in Mechanistic Studies

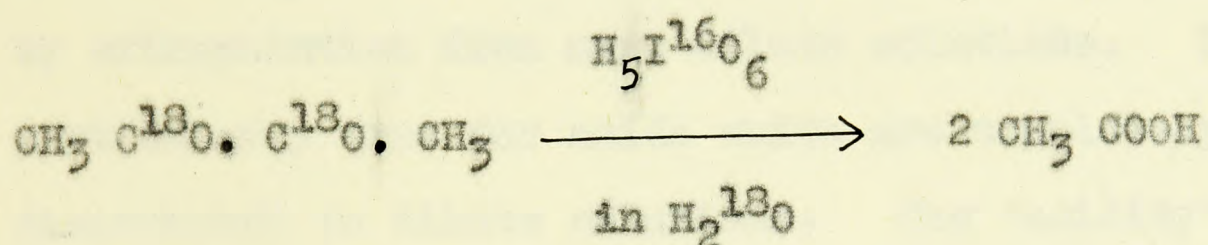
It is possible to prepare a wide variety of organic compounds containing isotopes carbon, nitrogen and oxygen.

^{13}C ^{14}C ^{15}N ^{18}O are the usual isotopes which have been used; they are available in the forms $\text{Ba}^{13}\text{CO}_3$, K^{13}CN , $\text{Ba}^{14}\text{CO}_3$, $^{15}\text{NH}_3$ and H_2^{18}O for the preparation of isotopically labelled compounds by normal chemical methods⁵⁰. The effect on reaction kinetics of isotopic substitution of C, N and O is much smaller than the effect of substituting hydrogen with deuterium and is of limited use in mechanistic determinations. The principal use of carbon, nitrogen and oxygen (and other) isotopes is in following the movement of a particular atom during a chemical reaction. The amount of the isotope in any particular product of a reaction can be determined with the use of a mass spectrometer. In the case of radioactive isotopes such as ^{14}C , the distribution of the isotope among the products can be found by measurements of their radioactivity. Two examples from the oxidation field are given to illustrate this use of isotopes in elucidating reaction mechanisms.

Bunton and Shiner³¹ oxidised pinacol with ^{18}O substituted periodic acid $\text{H}_5^{18}\text{O}_6$. The reaction proceeds through an intermediate complex of the form

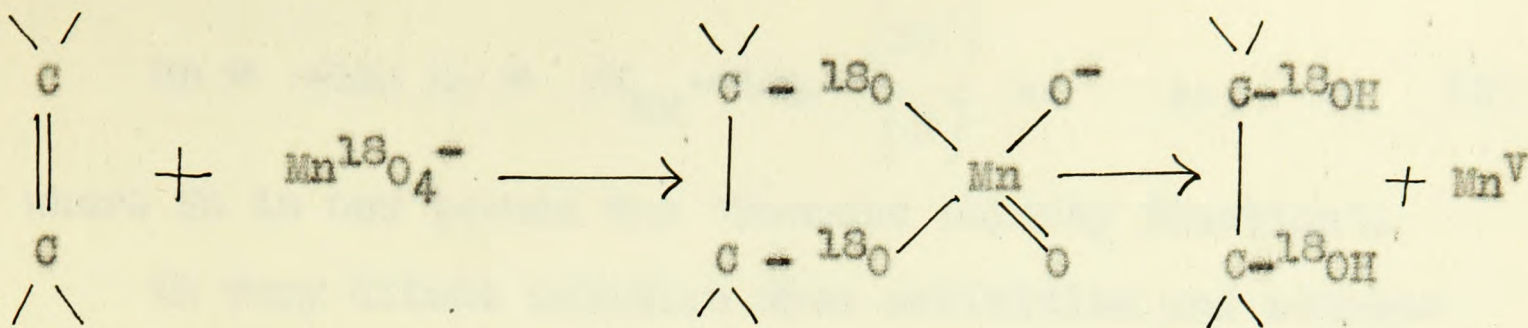


where the oxygen atoms a and b may come from the glycol or from the periodic acid. In the latter case the acetone produced in the oxidation will be largely $(\text{CH}_3)_2 \text{C}^{18}\text{O}$. In fact it was found to be almost entirely $(\text{CH}_3)_2 \text{C}^{16}\text{O}$, showing that the C - O bonds of the glycol remain intact during complex formation. In the oxidation of ^{18}O substituted biacetyl by ordinary periodic acid



the acetic acid, which was formed, contained about 50% of its oxygen as ^{18}O , after allowance had been made for exchange of isotopes between the periodic acid and water. This result was taken as evidence for a direct nucleophilic attack by periodate on the carbonyl group rather than for reaction through a hydrated form of the biacetyl.

Wiberg and Saegebarth³² oxidised oleic acid with ^{18}O substituted permanganate. They found that the oxygen in the resulting glycol was largely ^{18}O and took this to confirm a proposal that a cyclic intermediate was involved in the oxidation:-



Hammett's Acidity Function^{33,34}

The properties of concentrated solutions of acids often deviate considerably from what would be predicted by extrapolation from more dilute solutions. This is particularly true for acids which are completely dissociated in dilute solutions. The "acidity" of the solution towards indicators increases much more rapidly than its concentration.

Hammett and Deyrup³⁵ introduced an "acidity function" which is a measure of the tendency of a solution to protonate an uncharged base B. They defined a quantity h_0 by means of the equation

$$h_0 = K_{\text{BH}^+} + \frac{[\text{BH}^+]}{[\text{B}]} \quad \dots \quad (1)$$

where K_{BH^+} is the thermodynamic dissociation constant of BH^+ the protonated form of the base and terms in square brackets represent concentrations rather than activities. The logarithmic form of this equation is

$$H_0 = -\log h_0 = pK_{BH^+} - \log \frac{[BH^+]}{[B]} \quad \dots \quad (2)$$

where H_0 is now termed the "Hammett Acidity Function".

In very dilute solution when activities and concentrations become identical

$$K_{BH^+} = \frac{[B][H^+]}{[BH^+]}$$

so that $h_0 = [H^+]$ and $H_0 = pH$. In concentrated solutions h_0 is greater than $[H^+]$ by a factor which may be as much as 100. The variation of H_0 may be represented by inserting activity coefficients in equation (2).

$$H_0 = -\log [H^+] + \log \frac{f_{BH^+}}{f_B f_{H^+}} \quad \dots \quad (3)$$

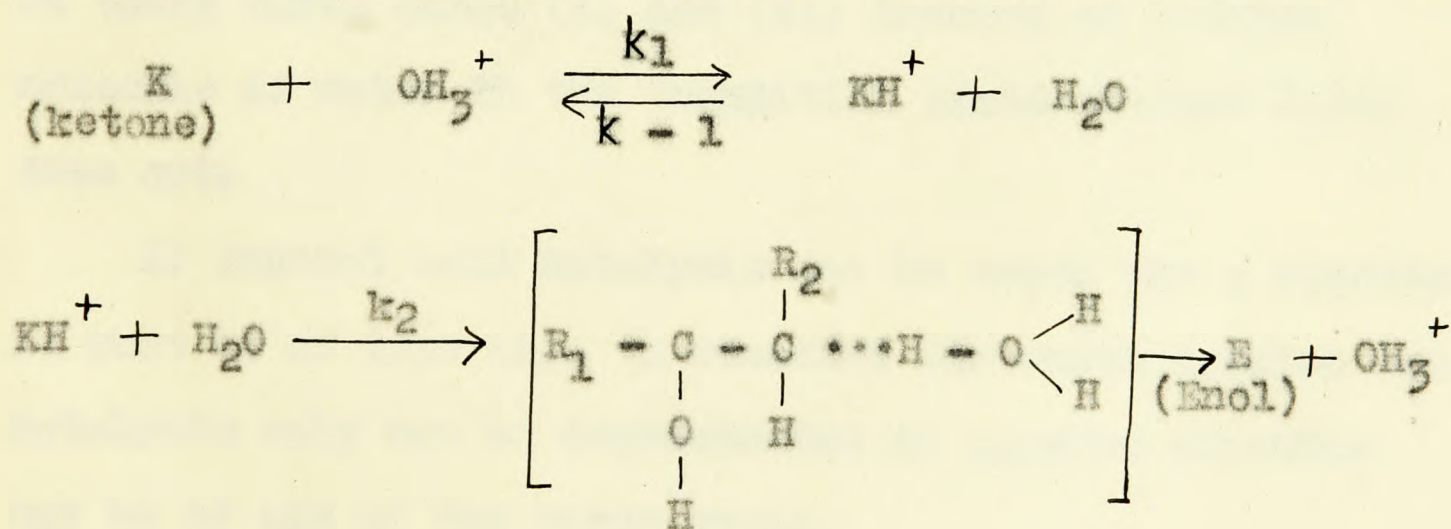
The acidity function as defined by equations (2) and (3) should be approximately independent of the base used, since values of $\frac{f_{BH^+}}{f_B}$ for various indicators will be nearly constant in dilute solutions of the indicator. This is born out by such experimental data as is available³⁶.

For aqueous solutions of many acids up to about 8M the acidity function is a well defined quantity and is approximately independent of the anion for several strong acids (e.g. HCl, HBr, H₂SO₄, HClO₄); above 8M discrepancies occur between the values for these acids and have been attributed to differences in the degree of hydration of

the protons.

The importance of the H_0 function in reaction kinetics is in the nature of the acidity dependence of acid catalysed reactions. Many reaction rates are accurately proportional to the concentration of hydrogen ions, but a wide variety of reactions have been reported which follow h_0 instead. This dependence on h_0 is in many cases only approximate but it is clear that there must be some significant difference in mechanism between these two classes of acid catalysed reaction in order to produce the observed acidity dependence.

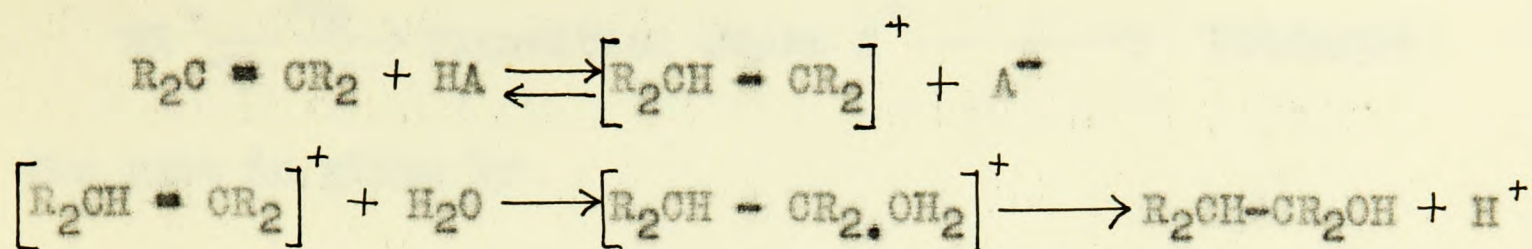
³⁷
Zucker and Hammett considered that there are three types of acid catalysed reaction (i) the rate determining step is a proton transfer as in enolisation



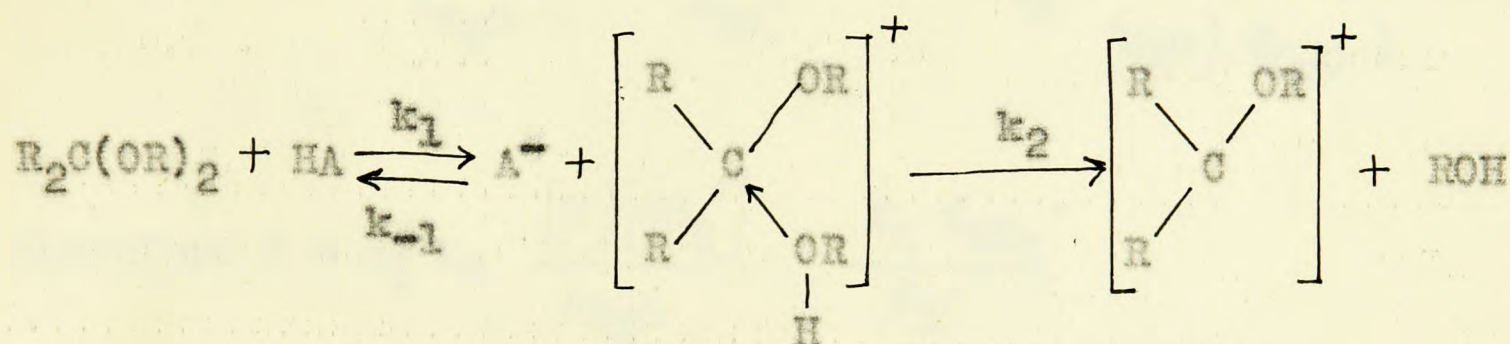
Transition State (X)

(ii) the rate determining step involves a water molecule, in some other way than as a base e.g. this mechanism for

hydration of an olefin



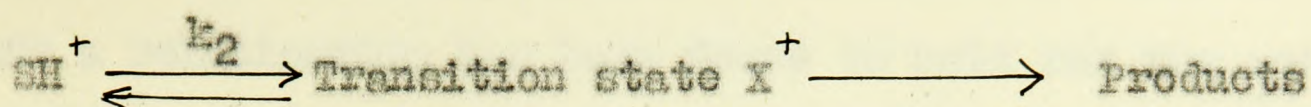
(iii) the rate determining step involves a spontaneous cleavage or other change of the conjugate acid of the substrate e.g. the following possible mechanism for hydrolysis of an acetal



Of these three cases (i) and (ii) involve an extra molecule of water in the transition state whereas (iii) does not.

If general acid catalysis can be shown for a reaction it must be of type (i). A reaction for which oxonium ion catalysis only can be demonstrated in aqueous solution may be of any of the three types.

For a reaction of type (iii), if the rate determining step is a first order reaction of the conjugate acid SH^+ of the substrate S



the rate is given by

$$V = k_2 [SH^+] \frac{f_{SH^+}}{f_{X^+}} \dots \dots \dots (4)$$

$$\text{Now } [SH^+] = K_a \frac{[S][OH_3^+]}{A_{H_2O}} \cdot \frac{f_S f_{OH_3^+}}{f_{SH^+}} \quad (A_{H_2O} = \text{Activity of water} = [H_2O] \cdot f_{H_2O})$$

$$\text{Therefore } V = k_2 K_a \frac{[S][OH_3^+]}{A_{H_2O}} \cdot \frac{f_S f_{OH_3^+}}{f_{X^+}}$$

$$\text{The rate constant } k = \frac{V}{[S]} = k_2 K_a [OH_3^+] \frac{f_S f_{OH_3^+}}{f_{X^+} A_{H_2O}}$$

$$\text{Substituting } h_0 = \frac{A_{H^+} f_B}{f_{BH^+}} \text{ from equation (1)}$$

$$k = k_2 K_a h_0 \frac{f_S f_{BH^+}}{f_{X^+} f_B}$$

It is not unreasonable to suppose that the activity coefficient of the transition state f_{X^+} bears a similar ratio to that of the substrate f_S as f_{BH^+} does to f_B . In this event the term $\frac{f_S f_{BH^+}}{f_{X^+} f_B}$ is constant and the rate k

becomes directly proportional to the acidity function h_0 . Such a relationship has been found to hold for the acid catalysed hydrolysis of sucrose and the nitric acid hydrolysis of cyanamide⁷⁵.

In the case of the enolisation type of reaction (1) above, the rate equation is

$$v = k_2 [KH^+] A_{H_2O} \cdot \frac{f_{KH^+}}{f_{X^+}}$$

which differs from equation 4 by the term A_{H_2O} .

$$\text{Substituting for } [KH^+] \text{ from } k_1 = \frac{A_{KH^+} A_{H_2O}}{A_K A_{OH_3^+}}$$

$$v = k_2 k_1 [K] A_{OH_3^+} \cdot \frac{f_K}{f_{X^+}}$$

$$\therefore \text{ The rate constant } k = k_2 k_1 [OH_3^+] \cdot \frac{f_K f_{OH_3^+}}{f_{X^+}}$$

Since the transition state X^+ contains one molecule of ketone, one molecule of water and a proton, the term

$\frac{f_K f_{OH_3^+}}{f_{X^+}}$ may well be approximately constant in which

case k becomes directly proportional to hydrogen ion concentration $[\text{OH}_3^+]$. This relationship was found to hold over a wide range of acid concentrations for the enolisation of acetophenone³⁷.

Thus, to summarize Hammett's views, acid catalysed reactions in which a water molecule is involved follow $[\text{H}^+]$ and not h_O , while acid catalysed reactions in which a water molecule is not involved follow h_O .

Since this postulation was made, many papers have been published both confirming the hypothesis and using it as a proved theory, but the hypothesis is really little more than an empirical relationship and should not be taken as definite proof of any particular reaction mechanism³⁸. In any case it only strictly applies to dilute aqueous solutions. Recently some papers have appeared claiming that the Zucker-Hammett hypothesis³⁹ is inadequate in certain instances and that whereas all reactions of type (iii) above should follow h_O , some reactions of type (ii) may follow h_O as well.

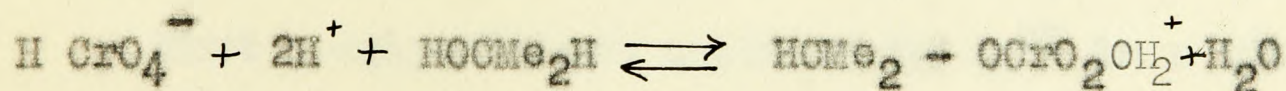
Established Mechanisms of Oxidation

In order to provide some basis for comparison of the reaction kinetics of V^V oxidations with other similar systems, a brief review of some reasonably well authenticated mechanisms of oxidation of organic compounds by inorganic species in aqueous solutions is given here.

1. Cr^{VI} oxidations^{45,46}

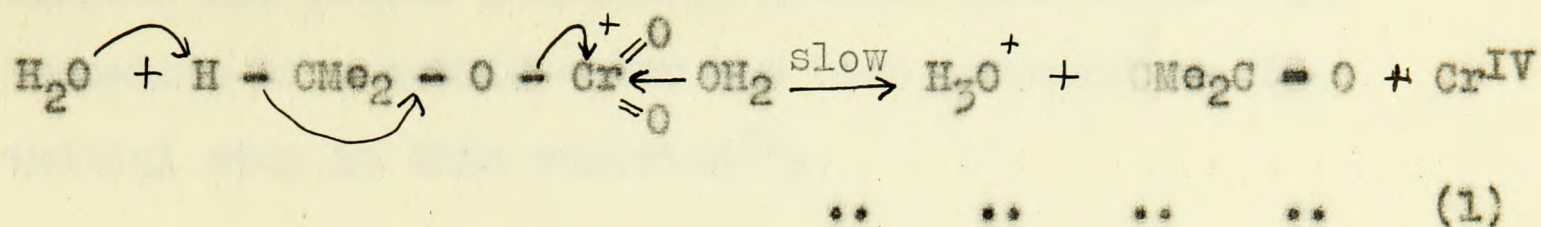
All Cr^{VI} oxidations must proceed by a stepwise reduction of the chromate ion, since a direct three electron transfer $Cr^{VI} \rightarrow Cr^{III}$ is impossible. The initial step in all cases may be a one electron $Cr^{VI} \rightarrow Cr^V$ or two electron $Cr^{VI} \rightarrow Cr^{IV}$ process. The intermediate Cr^{IV} and Cr^V species cannot be isolated and it is difficult to detect radical intermediates.

The Cr^{VI} oxidation of isopropyl alcohol is a reaction which has been studied very widely but a single mechanism for even the first stage of the reaction cannot be agreed upon by the various workers. It seems most likely that formation of a chromate ester is the initial step⁴⁰.

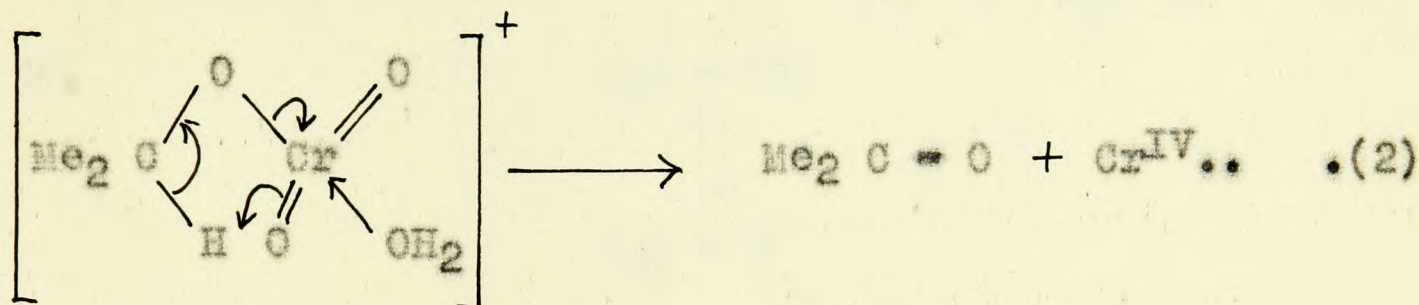


Such esters have in fact been isolated⁴⁰. They are unstable and may decompose either by base catalysis (which

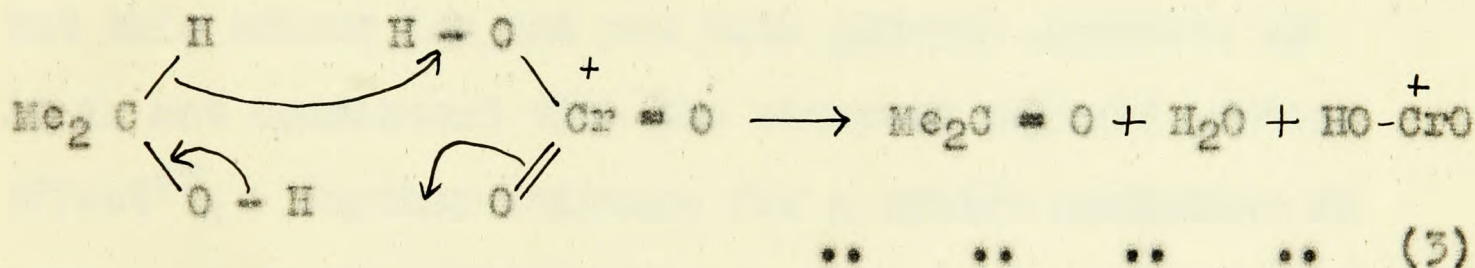
is not observed^{41,42}.)



or directly by a cyclic mechanism⁴³



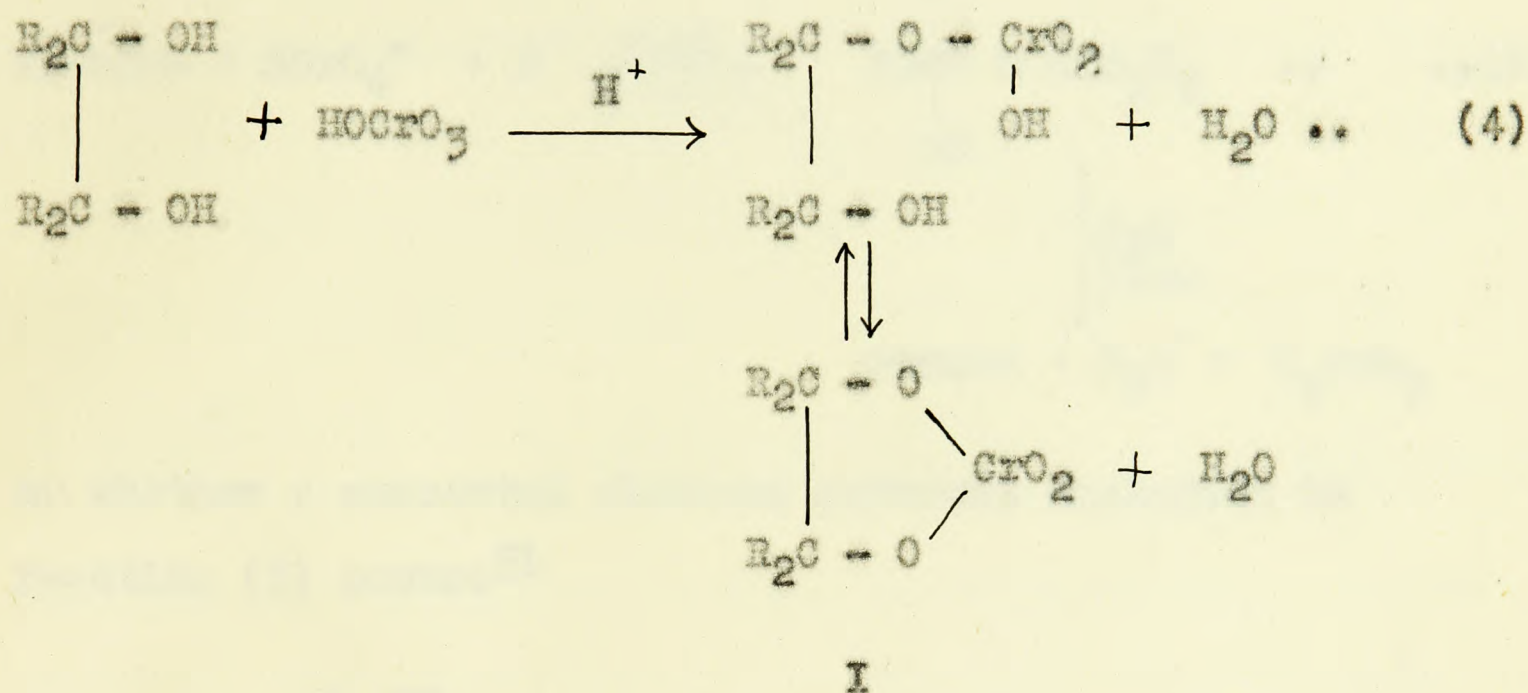
A third mechanism proposed⁴¹ does not involve ester formation but is inconsistent with the observed solvent isotope effect.



Diisopropyl ether is oxidised at only $\frac{1}{1500}$ of the rate at which isopropyl alcohol is oxidised, and it is significant that a chromate ester cannot be formed by this compound⁴⁴.

Most α -glycols are oxidised by Cr^{VI} in a similar manner to monohydric alcohols^{47,48}, but the fully substituted glycol pinacol is oxidised by a different mechanism, the rate of reaction being of first order with respect to $[\text{H}^+]$ instead of second order. The reaction involves C - C bond

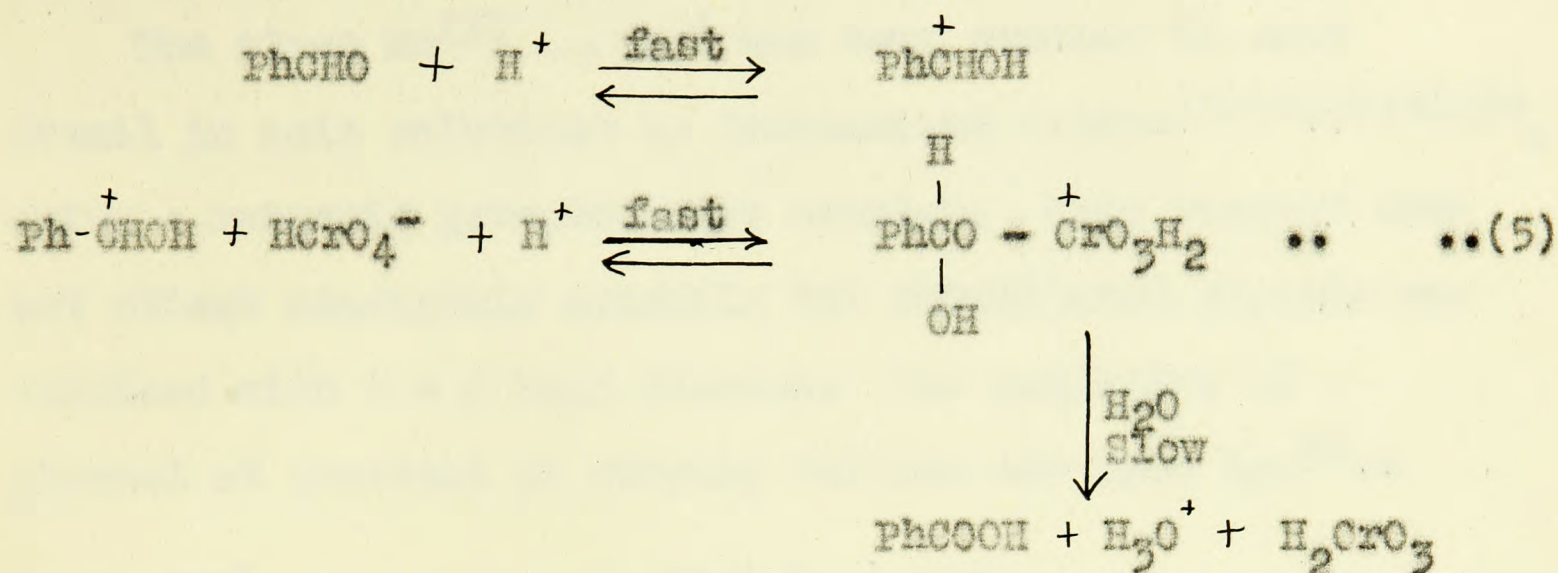
fission and yields 2 moles of acetone quantitatively. Formation of a cyclic ester has been proposed as the initial step in this reaction⁴⁸,



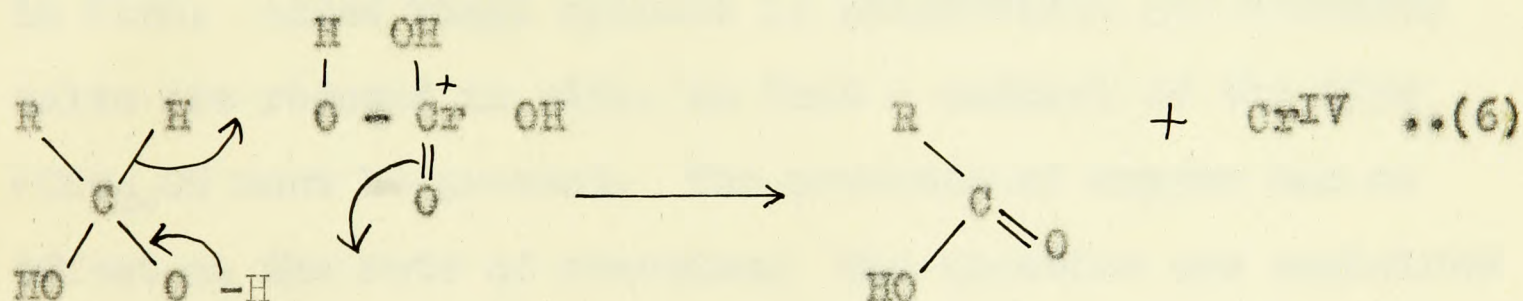
but this scheme has not met with general approval as it is not consistent with the observed solvent isotope effect⁴⁷. Further evidence for a cyclic mechanism of some sort is provided by the fact that pinacol monomethyl ether reacts at only $\frac{1}{1500}$ of the rate of pinacol. Presumably the ether is incapable of forming a cyclic complex such as I.

Studies of chromic acid oxidation of aldehydes have been concerned largely with benzaldehyde^{49,50,51} which can only give benzoic acid and is incapable of forming an enol or of giving C - C bond fission. Again there is argument as to whether the reaction proceeds through an ester such

as the half ester of benzaldehyde hydrate⁵⁰.



or whether a concerted electron movement analogous to reaction (3) occurs⁵¹



This latter mechanism is not consistent with the observed solvent isotope effect⁵².

2. Manganese

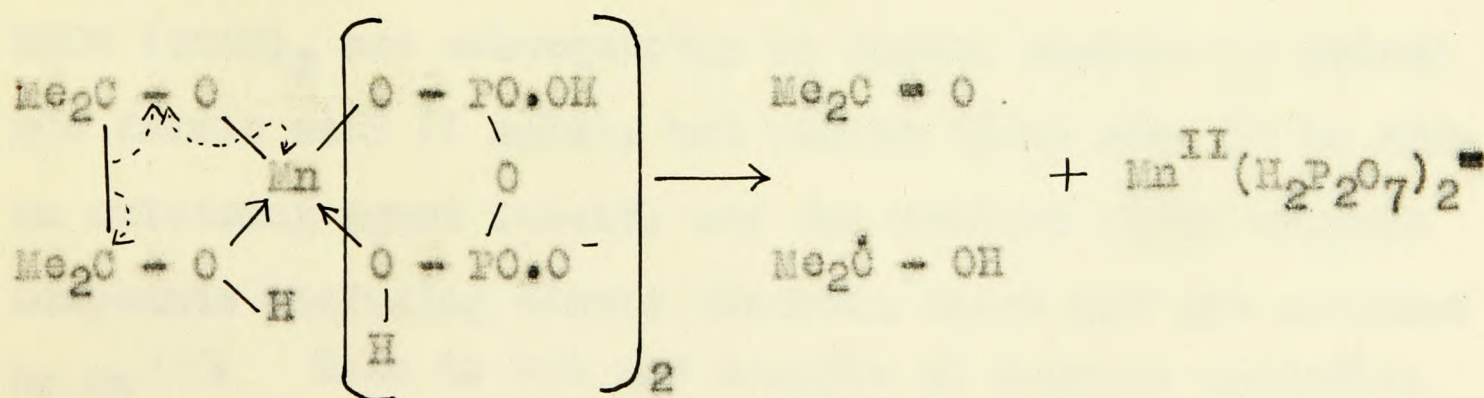
The wide variety of compounds oxidised by permanganate is well known, but the elucidation of a mechanism for any one oxidation is frustrated by the complex nature of the reaction. Since the change $\text{Mn}^{\text{VII}} \longrightarrow \text{Mn}^{\text{II}}$ in acid

solution involves the transfer of 5 electrons, at least 3 steps must be involved in all cases, probably more.

The stage $\text{Mn}^{\text{III}} \rightarrow \text{Mn}^{\text{II}}$ has been studied in some detail in acid solutions by Drummond and Waters^{15,18,20,53,54}, using a manganic pyrophosphate complex. This reagent does not attack monohydric alcohols but substituted glycols are oxidised with C - C bond fission. The oxidation of pinacol at constant pH roughly follows the rate law⁵³:-

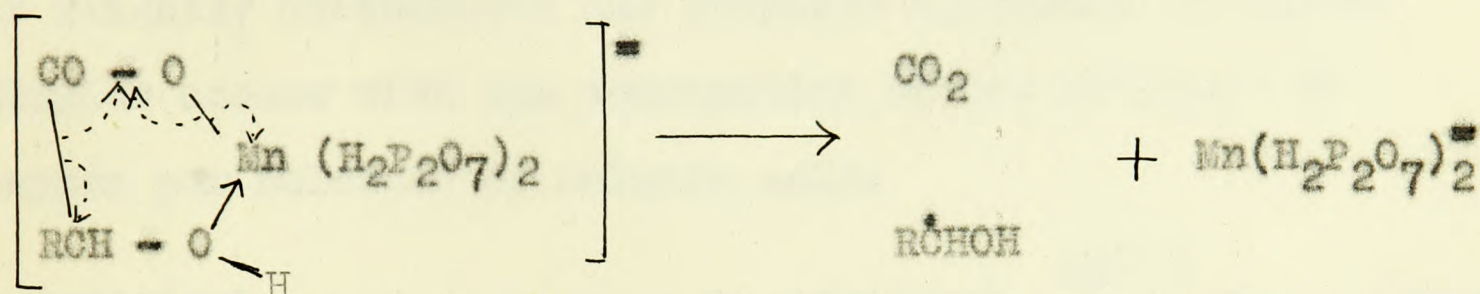
$$-\frac{d[\text{Mn}^{\text{III}}]}{dt} = \frac{A[\text{Mn}^{\text{III}}][\text{Pinacol}]}{b[\text{Pinacol}]} \cdot \frac{1}{[\text{Free pyrophosphate}]}$$

The dependence of rate on acidity is complex, being sigmoid in form. Added vinyl cyanide is polymerised and mercuric salts are reduced in situ, so that a radical of the type $\cdot\text{CMe}_2\text{OH}$ must be present. The presence of oxygen has no effect on the rate of reaction. The kinetics are explained by assuming that one molecule of pinacol replaces a pyrophosphate ion in the Mn^{III} complex and that the resulting complex spontaneously decomposes.



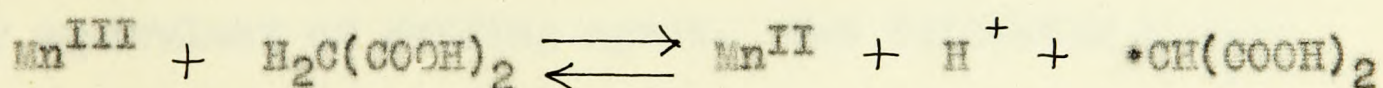
The oxidation of α -hydroxy acids⁵⁵ proceeds in a similar manner, the only difference being that manganous salts retard the oxidation.

The decomposition has been represented as



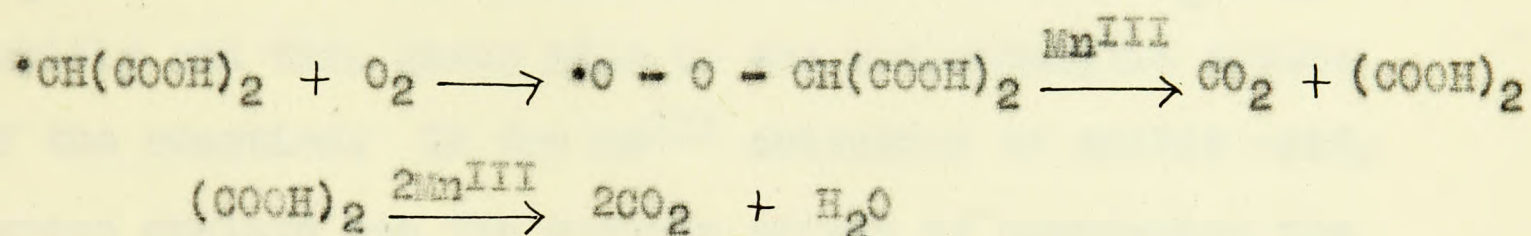
The radicals $\text{R}\dot{\text{C}}\text{HOH}$, or $\text{Me}_2\dot{\text{C}}\text{OH}$ from pinacol, are rapidly oxidised further. Lactic, malic and tartaric acids are all oxidised quantitatively to carbon dioxide, water and one mole of formic acid.

The oxidation of malonic acid⁵⁴ also proceeds through a chelate complex. The first stage again involves the formation of a free radical $\cdot\text{CH}(\text{COOH})_2$

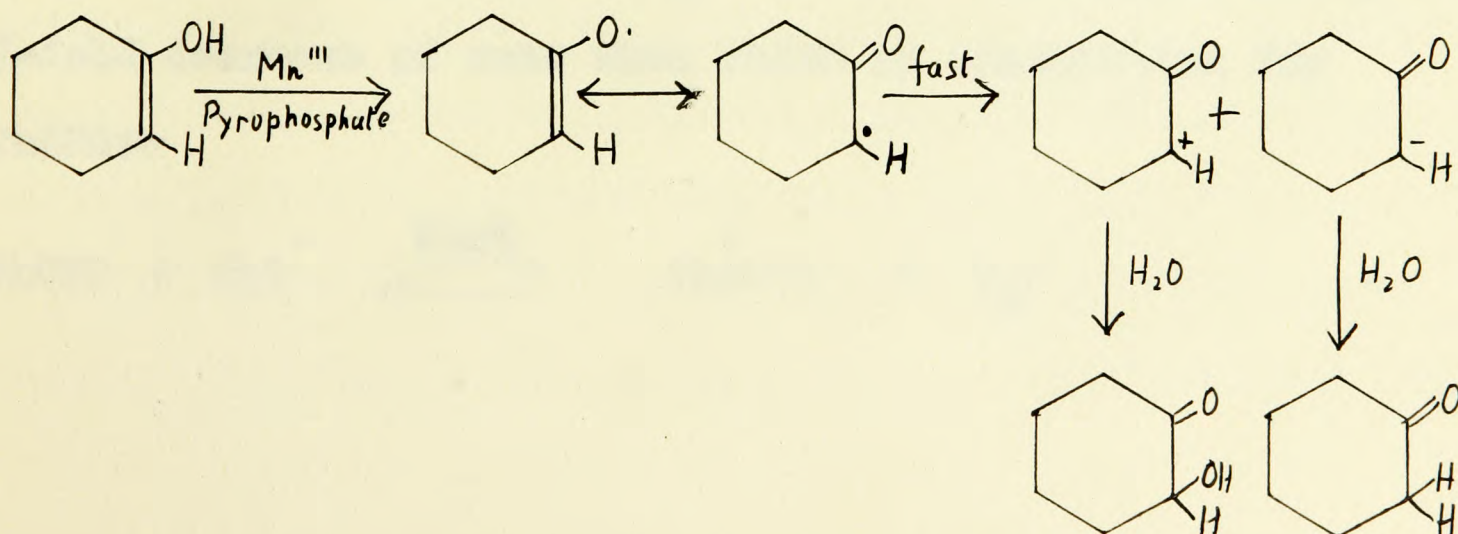


This radical can readily be oxidised to tartronic acid $\text{CHOH}(\text{COOH})_2$ and subsequently to carbon dioxide (2 moles) and formic acid (1 mole), but unlike $\text{R}\dot{\text{C}}\text{HOH}$ etc it is also an oxidising agent itself, and can oxidise other organic compounds including simple alcohols which are not attacked by Mn^{III} . This is but one example of induced oxidation,

a phenomenon which is possible for one-electron oxidations but not for two-electron oxidations. Oxygen also has a notable effect on the reaction, as it attacks the intermediate radical. In the presence of oxygen the reaction is slightly accelerated and complete oxidation to carbon dioxide occurs with the consumption of one molecule of oxygen per molecule of malonic acid.



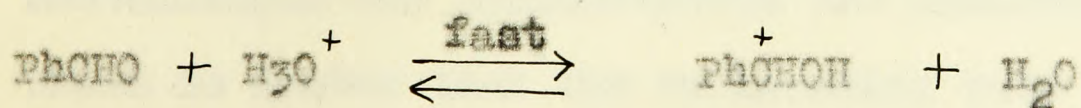
The Mn^{III} oxidation of a variety of ketones has been investigated¹⁸ and found to proceed by attack on the enolic form of the ketone. In particular the initial rate of oxidation of cyclohexanone increases with increasing Mn^{III} concentration to a limit equal to the rate of enolisation per equivalent of oxidizing agent. The following scheme was proposed by Drummond and Waters.

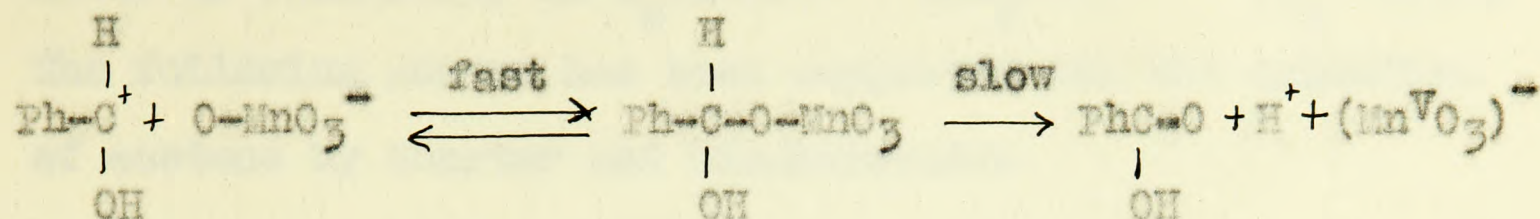


The ring system is then further degraded.

Among the few permanganate oxidations which have been thoroughly investigated is that of oxalic acid, which is catalysed by Mn^{2+} ions⁵⁶. The initial attack has been shown to be due to Mn^{3+} ions, which form a complex with oxalate groups, and not to MnO_4^- itself. The radical ion $\dot{\text{C}}\text{O}_2^-$ which is formed intermediately, attacks MnO_4^- very rapidly and this gives rise to the autocatalytic nature of the reaction. In the Mn^{III} oxidation of oxalic acid, excess oxalate has the curious effect of decreasing the reaction rate. This anomaly is attributed to the oxidation occurring by the decomposition of the 3 species $\text{Mn}(\text{C}_2\text{O}_4)^+$, $\text{Mn}(\text{C}_2\text{O}_4)_2^-$, $\text{Mn}(\text{C}_2\text{O}_4)_3^{3-}$, the first of these reacting fastest, but being decreased in abundance by excess oxalate⁵⁷.

Permanganate oxidation of aromatic aldehyde⁵⁸ has been assigned the following mechanism on the basis of observations which include general acid catalysis and a 7-fold decrease of rate when PhCDO is substituted for PhCHO :-



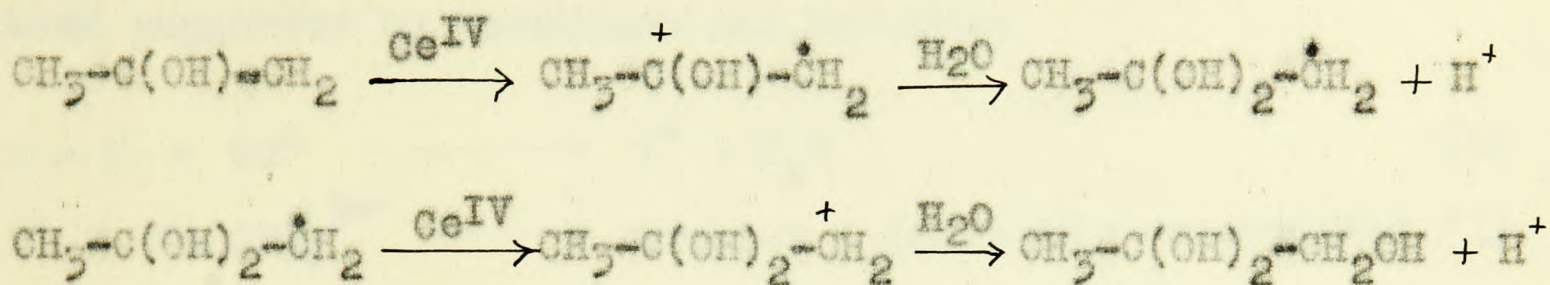


Isotopically labelled oxygen in the permanganate is transferred to the organic compound and can be found in the benzoic acid. The reaction is more complex in alkaline solution.

3. Cerium^{IV} 59,60

This reagent is a confirmed one-electron oxidising agent; its reactions involve free radical intermediates. It does not attack monohydric alcohols at low acidities (of Mn^{III}). The oxidation of aldehydes and ketones has been studied and a mechanism involving prior enolisation of the ketone has been suggested⁶⁰. The rate of reaction is of first order with respect to both ketone and Ce^{IV} at low concentrations of oxidising agent, but becomes independent of $[\text{Ce}^{\text{IV}}]$ at higher concentrations, the rate in most cases approaching that of enolisation. However acetaldehyde and cyclohexanone are oxidised at a rate which is faster than the enolisation rate and therefore must be attacked in the carbonyl form. In other cases the

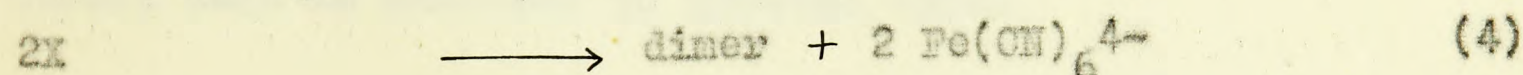
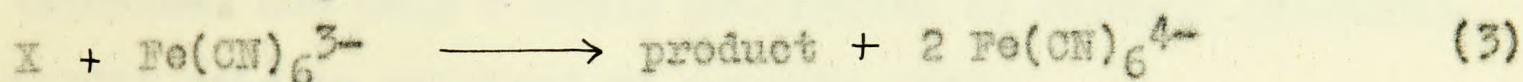
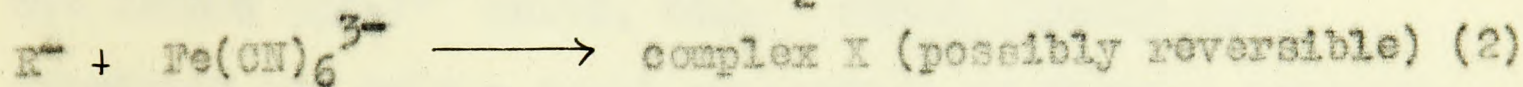
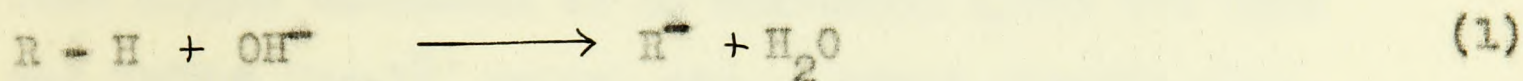
order of reactivity is $R_2CH-CO-$ RCH_2-CO- CH_3-CO- .
 The following scheme has been suggested for the oxidation of acetone by Shorter and Hinshelwood:-



4. $Fe(CN)_6^{3-}$ 61

Although this reagent gives some typical one-electron oxidations, it fails to induce vinyl polymerisation which is a typical reaction of free radicals. It is believed that this is because all radical products which are present are complexed to the ferricyanide and do not exist free. In the oxidation of some simple aldehydes, ketones and nitro compounds in alkaline solution, the initial rate of oxidation is of first order with respect to both alkali (except in the oxidation of isobutyraldehyde) and to substrate. The dependence on $[Fe(CN)_6^{3-}]$ is zero order with isobutyraldehyde, first order dropping to zero order at high concentrations for n-butanal and diethyl ketone, first order rising to second order for acetone and nitroethane, and first order throughout for

iso-propanal, ethyl methyl ketone, di-isopropyl ketone and 2-nitropropane. The aldehydes and ketones must react in their enol forms and the following scheme has been suggested by Speakmann and Waters:-

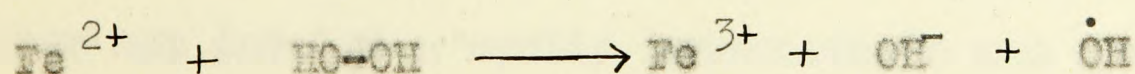


For iso-butanal (1) is slower than (2), the reverse applying to other compounds. For n-butanal and diethyl ketone (1) and (2) may have comparable rates. For acetone and nitroethane (2) may reach equilibrium and (3) become rate determining.

Benzaldehyde is not attacked by $Fe(CN)_6^{3-}$ and furfuraldehyde is only attacked very slowly. These compounds cannot enolise normally.

5. Fenton's Reagent^{16,17}

This is another one-electron oxidising agent. The reactive species is the hydroxyl radical formed by direct reaction of hydrogen peroxide with the ferrous ion.

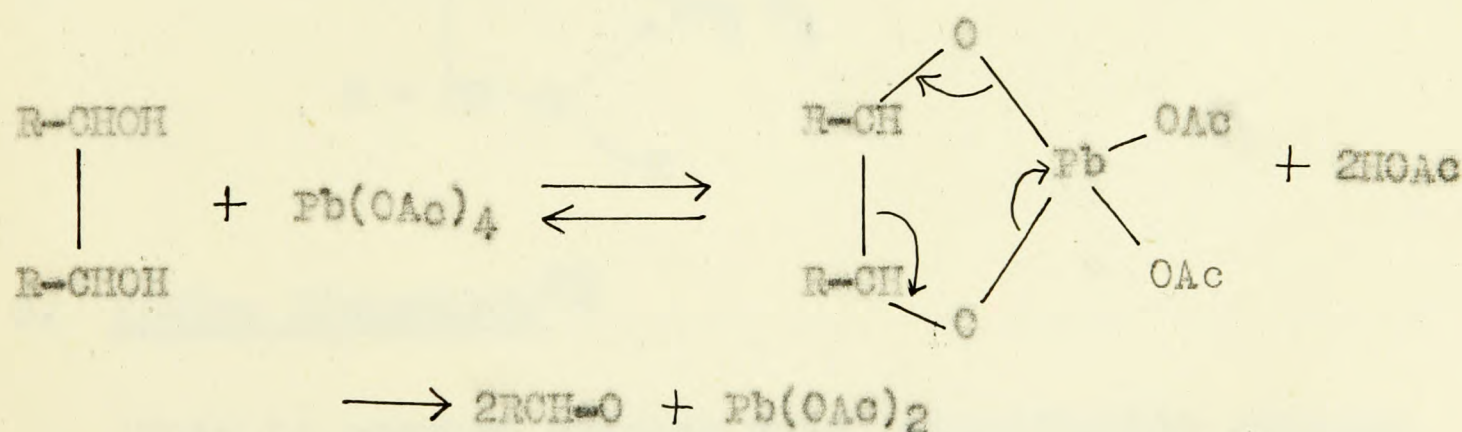


It has been suggested that pinacol is oxidised through the diradical $(\text{CH}_3)_2 - \underset{\underset{\cdot}{\text{O}}}{\underset{|}{\text{C}}} - \underset{\underset{\cdot}{\text{O}}}{\underset{|}{\text{C}}} - (\text{CH}_3)_2$

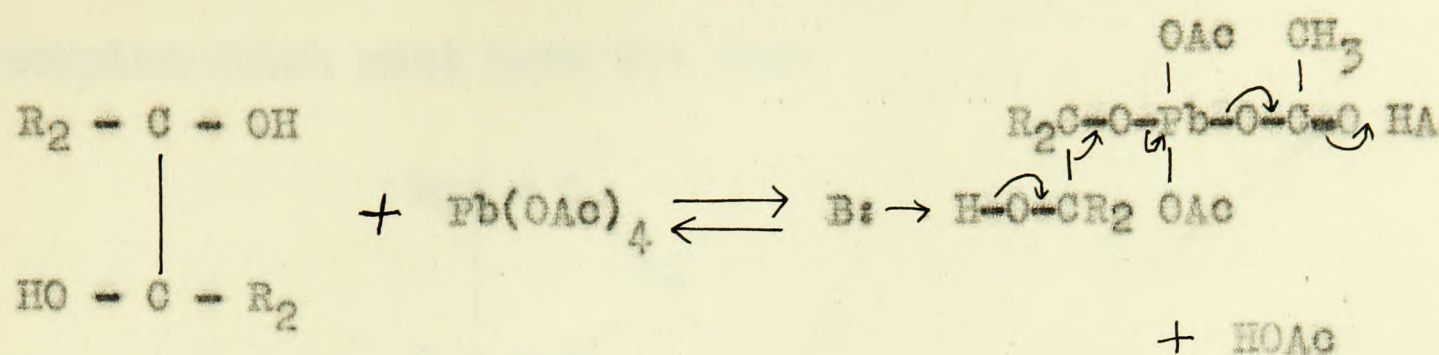
which undergoes C - C bond fission to give acetone. This oxygen diradical is an oxidising agent and does not reduce $(\text{Hg}_2)^{2+}$ salts, etc. Other glycols give diketones and only a small proportion of fission products. Lactic acid is oxidised to pyruvic acid.

6. Lead Tetraacetate⁶²

A great deal of attention has been focussed on this reagent which is well known for its specific glycol fission reaction. Various conflicting views have been advanced about the mechanism of the reaction, but the following scheme now seems to have been accepted.



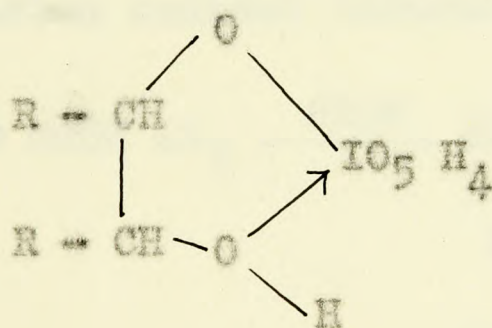
For trans-ditertiary glycols a slower mechanism which does not involve a cyclic intermediate was proposed by Bell, Rivlin and Waters.⁶³



(B = Base, HA = acid)

7. Periodic Acid⁶⁴

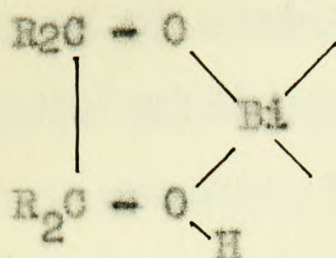
This reagent is also specific for glycol fission, but unlike lead tetraacetate it cannot oxidise trans glycols with the hydroxy groups held rigidly 180° apart as in 9.10 trans decalin diol. The mechanism of the reaction must involve a cyclic complex such as



8. Sodium Bismuthate⁶⁵

This is another reagent which specifically attacks α -glycols, but it has not been so widely examined as other such reagents. Reaction again occurs through a cyclic

complex which must have the form

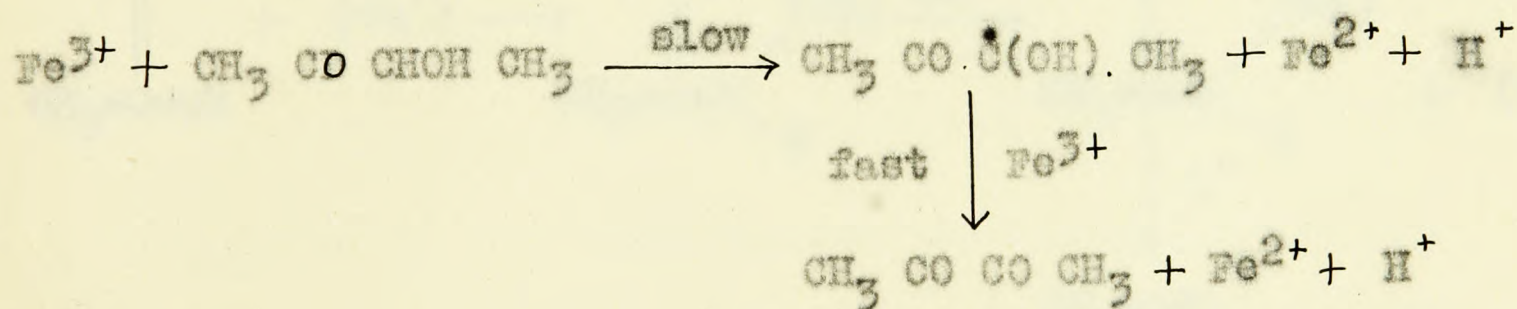


α -Glycols give aldehydes or ketones on oxidation; acyloins give a carbonyl compound and a carboxylic acid.

These three glycol splitting reagents are typical two-electron oxidising agents.

9. Ferric Salts

The reaction of ferric salts with acetoin has been studied⁶⁶ and shown to proceed by a series of one-electron steps with carbon radical intermediates.

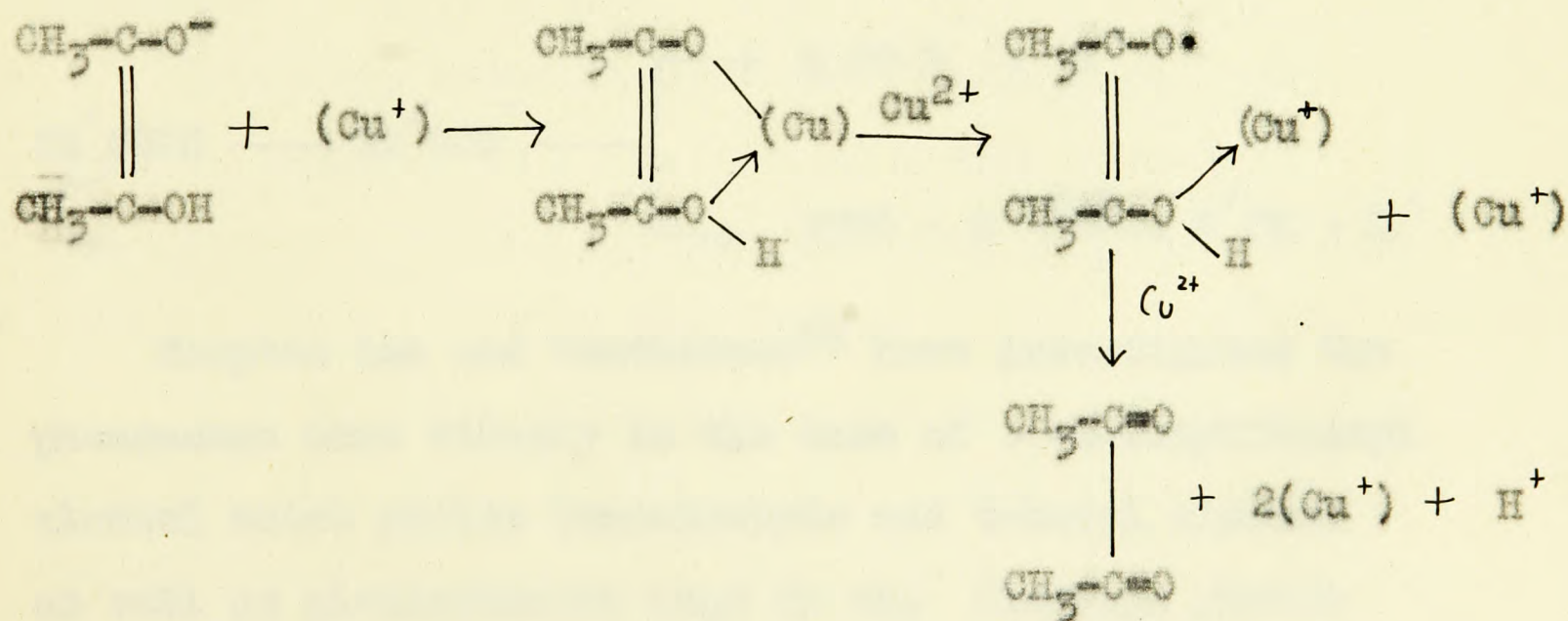


The reaction induces polymerisation of methyl methacrylate and is partially suppressed by oxygen.

The methyl substituted acetoin $\text{CH}_3 \text{ C}(\text{CH}_3)(\text{OH}) \text{ CO CH}_3$ is not attacked by this reagent.

10. Complexed Cupric Salts⁶⁷

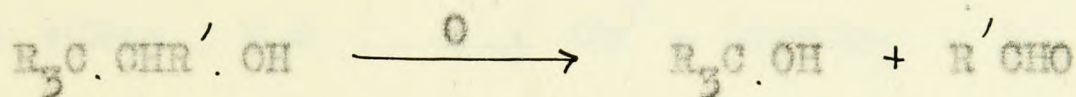
The oxidations of acetoin and glucose by alkaline solutions of cupric salts complexed with picolinic acid or citric acid are preceded by an induction period and thereafter are substantially of zero order with respect to $[\text{Cu}^{2+}]$ and first order with respect to organic substrate. The induction period is due to reoxidation by dissolved oxygen of the cuprous complex which is formed, and can be eliminated by excluding oxygen from the system. Oxidation must occur through the ene-diol which forms a chelated cuprous complex that is easily attacked by cupric ions to give biacetyl and 2 equivalents of cuprous oxide.



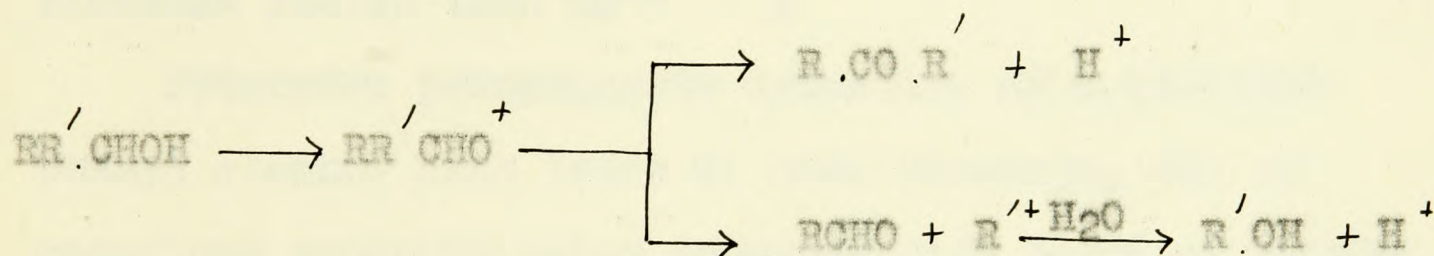
This system is similar to the ferrioxanide oxidations in that all intermediate radicals are complexed and cannot be detected in the normal manner.

CHAPTER 3Mechanisms of Oxidation of Some Secondary Alcohols

Mosher and his colleagues⁶⁸ have drawn attention to the fact that several secondary alcohols, particularly those of the type $R_3C \cdot CHR' \cdot OH$, suffer carbon-carbon bond fission when oxidised by chromic acid, in addition to oxidation to the corresponding ketone.

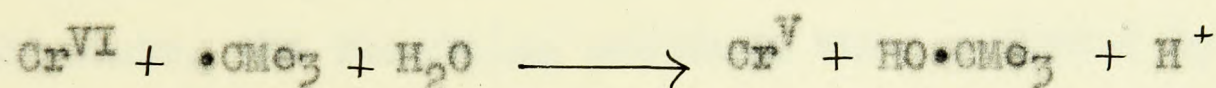
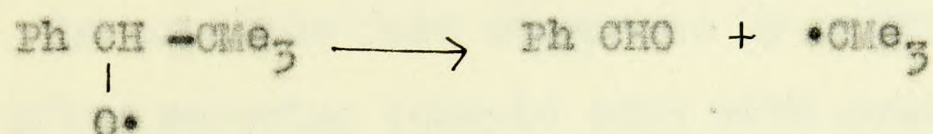
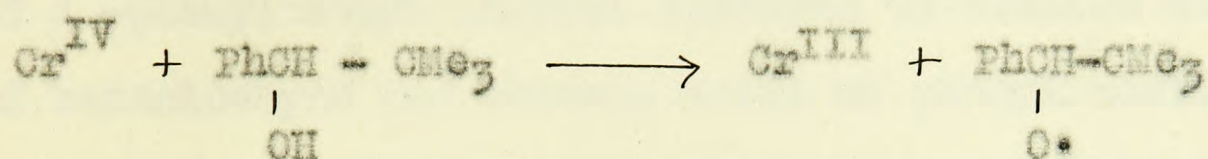


They attributed this anomaly to transient oxy cation formation.



Hampton Lee and Westheimer⁶⁹ have investigated the phenomenon more closely in the case of α -t-butylbenzyl alcohol which yields benzaldehyde and t-butyl alcohol as well as pivalophenone $Me_3C \cdot CO \cdot Ph$. Cleavage yields up to 70% were obtained, but this figure was sharply decreased in the presence of Mn^{2+} or Ce^{3+} ions which remove

any Cr^{IV} intermediate as soon as it is formed. It was therefore suggested that the cleavage was caused by Cr^{V} (formed from $\text{Cr}^{\text{IV}} + \text{Cr}^{\text{VI}}$) or possibly by a one electron process by Cr^{IV} e.g.:-



It has recently been shown that the salt trisodium chromate, which is known to contain Cr^{V} , gives greater cleavage yields than Cr^{VI} 70.

Potassium permanganate oxidation of α -*t*-butylbenzyl alcohol also leads to some cleavage, 46% of anomalous products being reported with 34.3% of ketone, while α -*t*-amylbenzyl alcohol gave 56% cleavage 71.

In the case of the primary alcohol β -phenylethyl alcohol, chromic acid oxidation has also been reported to give some benzaldehyde. This may be due to a similar direct C - C bond cleavage of the alcohol or, possibly, oxidation through β -phenylacetaldehyde.

It was therefore decided to investigate the products

of the oxidations of α -t-butylbenzyl alcohol and β -phenylethyl alcohol by the known one-electron oxidising agent V^V . High yields of cleavage products were obtained with both alcohols. The non-volatile products of oxidation of β -phenylethyl alcohol appeared to consist entirely of benzaldehyde and benzoic acid; no phenylacetic acid could be detected. Formic acid was found among the more volatile products. The products were separated by extraction into ether and after removing benzoic acid with aqueous sodium carbonate, the remaining products were analysed by gas-liquid chromatography. The benzoic acid was isolated on acidification of the sodium carbonate extract and had M.Pt. and mixed M.Pt. 121° which showed also the absence of other non volatile acids. The aqueous solution however gave a strong chromotropic acid test for formic acid. Table I gives the yields of products obtained when 2g. portions of alcohol were oxidised at 100° .

Table I

$H_2SO_4(M)$	Reaction Time (Hr)	Products detected (mole. %)		
		$PhCH_2CH_2OH$	$PhCHO$	$PhCOOH$
0.7	24	4	36	38
2.5	4	0	74	18
3.4	4	0	64	24

Under the conditions of these reactions in 3.4M sulphuric acid, pure benzaldehyde was found to be oxidised to benzoic acid to the extent of 20% under nitrogen, or 30% in air during 4 hours. Thus the above results are consistent with the theory that all the benzoic acid arises from oxidation of benzaldehyde. No reduction of V^V in 3.4M sulphuric acid occurred during 8 hr. in the presence of phenylacetic acid, but oxidation of benzyl alcohol to benzaldehyde was complete in a few minutes. Dehydration of β -phenyl-ethyl alcohol to styrene could not be detected after 8 hr. in 4.5M sulphuric acid at 100° under nitrogen.

Table 2 gives the yields of the various products obtained in the V^V oxidation of α -t-butyl ^{benzyl}~~phenyl~~ alcohol at 100° . An olefin was found to be formed by dehydration of the alcohol in sulphuric acid. This was assumed to be Ph CMe: CMe₂ which is reported by various authors⁷⁴ to be formed in similar circumstances and must be formed by a Wagner rearrangement. The material is oxidised by permanganate to acetone. Experiment no 4 in table 2 was performed with alcohol which had been treated for 16 hr. at 100° with 5M sulphuric acid to convert it into the olefin. Evidently this is not oxidised to benzaldehyde

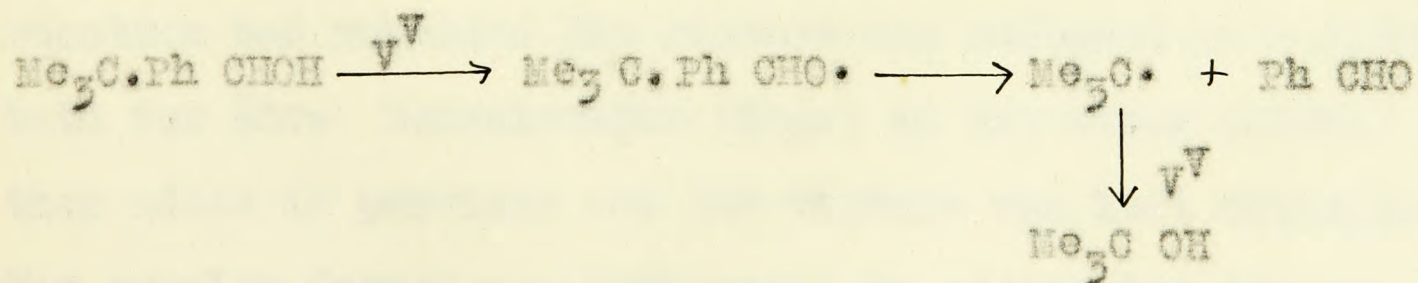
Table 2. Oxidation of α -t-butylbenzyl alcohol by vanadium (V).

Expt. no.	Alcohol (g.)	NH_4VO_3 (g.)	H_2SO_4 (M)	Total Reaction		Products found (moles %)					
				vol. (ml.)	time (hr.)	Alcohol	Ph. CHO	BzOH	Ph. CO. Bu ^t	Ph. CMe: CMe ₂	
1a	2.0	20	1.1	200	10	100	-	-	-	-	-
b	2.0	20	1.8	220	10	>95	-	-	-	-	-
c	2.0	20	2.4	240	10	5	39	24	15	6	
2	1.75	15	3.4	200	4	10	36	24	15	13	
3	2.0	15	4.5	200	4	0	32	48	10	10	
4	0.5	10	4.5	150	3	0	0	70	-	-	

by V^V . Acetophenone which might be expected as the first oxidation product was found to be rapidly and completely oxidised to benzoic acid under the conditions used. The true cleavage yields in the oxidation are thus represented by the yield of benzaldehyde together with a proportion of the yield of benzoic acid, and have a maximum of about 50%.

The oxidation of benzhydrol by V^V was found not to give any benzoic acid and the only product which could be detected was benzophenone (dinitrophenylhydrazone M.Pt. and mixed M.Pt. 235-237°).

The products of oxidation of β -phenyl-ethyl alcohol are consistent either with direct C - C bond fission of the alcohol, or with oxidation through the aldehyde, by the mechanism to be discussed later. In the case of α -t-butyl-benzyl alcohol the results establish that carbon-carbon bond fission can occur directly when alcohols other than 1,2-glycols are oxidised by one-electron abstracting agents and are consistent with the mechanism



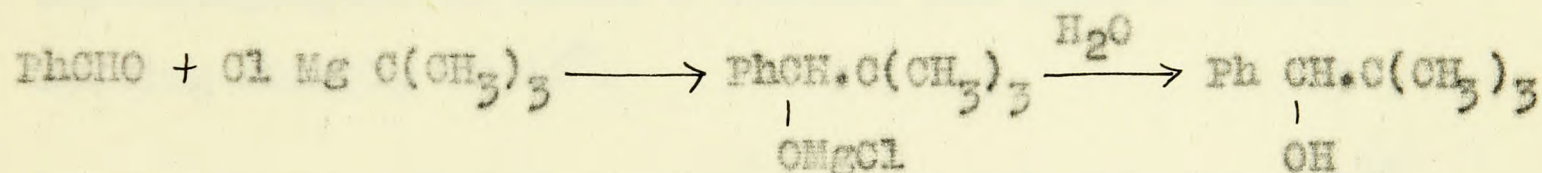
By analogy with this, the anomalous chromic acid

oxidation of α -t-butyl^{benzyl}~~benzyl~~ alcohol is quite likely to proceed by a one-electron oxidation by Cr^{IV} as shown on p 43

A single oxidation of α -t-butylbenzyl alcohol with ceric sulphate was conducted under similar conditions to the V^{V} oxidations. The same products were obtained, but the gas-liquid chromatogram gave an extra peak which could not be identified with any suspected product.

Experimental

Preparation of α -t-butylbenzyl alcohol ⁶⁹



t-Butyl chloride (148g.) was added in portions to magnesium turnings (36g.) and ether (300ml.) dried over phosphorus pentoxide. A crystal of iodine and a drop of methyl iodide were added to facilitate reaction. When the reaction had subsided the mixture was refluxed on a water bath for ½ hr. Benzaldehyde (50g.) in dry ether (50ml.) was then added in portions and the mixture was left overnight. The complex formed was decomposed by adding ice (100g.) and concentrated hydrochloric acid (150ml.) The ether layer was separated, and together with further ether extracts

of the aqueous layer was washed successively with aqueous sodium carbonate solution, aqueous sodium bisulphite and water and finally dried. After evaporation of the ether the residue was distilled at 120° - 140° /30-35mm. yielding 50g. impure solid α -t-butylbenzyl alcohol. The material was redistilled at 55 - 60° /0.1mm. and gave a practically colourless solid of melting point 35 - 38° , rising to 42 - 44° after it had been kept for a month. This material gave no positive test for ketonic substances and was used in some preliminary experiments. For the quantitative oxidations, the crude product from a separate preparation was chromatographed on Grade H alumina and gave a fraction melting sharply at 43° .

Preparation of Pivalophenone (Phenyl t-butyl ketone)⁷²



A sample of pivaloyl chloride which had been prepared from pivalic acid was kindly supplied by Dr. M. Clubb. Phenyl magnesium bromide prepared from bromobenzene (16.3g) and magnesium turnings (2.52g.) in dry ether (150ml.) was added during a period of $2\frac{1}{2}$ hrs. to a stirred solution of pivaloyl chloride (7.9g.) in dry ether (200ml.), which had been cooled in an ice-salt mixture to below 0° . The

reaction mixture was left overnight and was then decomposed with dilute hydrochloric acid to give the free ketone which was extracted into ether, washed with aqueous sodium bicarbonate, aqueous sodium bisulphite and water and then dried. After evaporation of the ether the residue was distilled at $100-140^{\circ}/40\text{mm.}$ and redistilled at $65-70^{\circ}/0.2\text{mm.}$ This product was chromatographed on Grade H alumina and eventually gave a product distilling at $70-72^{\circ}/0.4\text{mm.}$; this was quite colourless. $n_D^{25} = 1.510$ (reported 1.5102^{72})
M.Pt of oxime 167° (reported 167°).

Preparation of t-butyl benzoate⁷³

Anhydrous t-butanol (20ml.), benzoyl chloride (16.8ml.) and pyridine (24.4cc.) were mixed and left to stand for 24hr. White crystals were deposited. The mixture was treated with 30 ml. water which caused an oil to separate. This oil was washed with aqueous sodium carbonate and water and was then dried and distilled at $116^{\circ}/23\text{ mm.}$; it had $n_D^{25} = 1.4879$. The ester was found to decompose on a gas-liquid chromatography column at 150° and could not be responsible for any of the peaks in the chromatogram obtained from the products of the alcohol oxidations.

Chlorobenzene

Chlorobenzene (200g.) was stirred for 4 hr. on a boiling water bath with chromic oxide (50g.) in 3.5N sulphuric acid (250ml.). The chlorobenzene was separated, washed with water, dried over calcium chloride and distilled from a little sodium carbonate at 131°.

Gas Liquid Chromatography

A 187cm. column of Apieson L grease (10% w/w) on embacel was used with a hydrogen flame ionisation detector. The column was operated at 150°. Under these conditions the materials used took up to 40 minutes to pass through the column, α -t-butyl benzyl alcohol being last to come off. Undecane and acetophenone were found to be satisfactory as standard compounds for estimating absolute amounts of products. Results obtained were consistent to within 2 or 3%.

Method of performing quantitative oxidations

Portions of the alcohol (generally 2g.), dissolved in purified chlorobenzene (10ml.) were stirred vigorously with solutions of ammonium metavanadate in aqueous sulphuric acid in a 500 ml. flask which was closed through a mercury

trap around the stirrer, and which was heated on a boiling water bath. In many cases the vanadium was present partly as a suspension of vanadium pentoxide but this dissolved as reaction proceeded to leave a clear blue solution of V^{IV} . The resulting mixture was extracted with ether, from which sodium carbonate solution removed any benzoic acid which had been formed. This acid was precipitated on acidification and was isolated and weighed, together with a small amount which was extracted with ether from the acidified solution. The main ether extract was dried with magnesium sulphate and, after evaporation of most of the ether, was analysed by gas-liquid chromatography. In some early experiments benzaldehyde was removed from the ether extract with sodium bisulphite solution. Pivalophenone was identified in the residue as its oxime M.Pt. and mixed M.Pt. 165-167°.

The above method of oxidation was also adopted for other materials insoluble or only slightly soluble in water. The use of a 3-necked flask enabled nitrogen to be bubbled through the reaction system if required and volatile materials formed during oxidation to be drawn off and analysed.

Dehydration of α -t-butylbenzyl alcohol to an olefin

The alcohol (2g.), dissolved in chlorobenzene (10ml.) was vigorously stirred for 16 hr. with 5M sulphuric acid (200ml.). After this treatment, investigation of the chlorobenzene layer by gas-liquid chromatography revealed that the alcohol had been completely converted into a different material. The chlorobenzene solution rapidly decolourised both warm dilute potassium permanganate solution and cold bromine water. The material gave no reactions characteristic of compounds containing oxygen functions and gave no infra red absorption in the regions characteristic of -OH and -CO groups. In particular 2,4 dinitrophenylhydrazine gave no derivative.

An acid solution of potassium permanganate was stirred in a 3-necked flask on a boiling water bath. A current of air was drawn out through one neck of the flask and was bubbled through a solution of 2,4 dinitrophenylhydrazine in aqueous methanol and hydrochloric acid. On adding a portion of the chlorobenzene solution of the olefin to the permanganate solution, the vapours evolved immediately gave a thick yellow precipitate with the 2,4 dinitrophenylhydrazine. This had a M.Pt. of 127° which was unchanged after admixture with authentic acetone 2,4 dinitrophenyl-

hydrazone.

Treatment of β -phenylethanol with strong acid in the same manner but under nitrogen did not give an olefinic compound. The infra red spectrum of the alcohol was the same both before and after this treatment.

CHAPTER 4Oxidation of Some α -Hydroxy Acids and of Pinacol

55

Levesley and Waters have studied the oxidation of some simple α -hydroxy acids with manganic pyrophosphate. They found that lactic, tartaric and malic acids are gradually oxidised to carbon dioxide and formic acid; no formaldehyde is formed. A study of the reaction kinetics showed that the oxidations fit a rate equation of the form

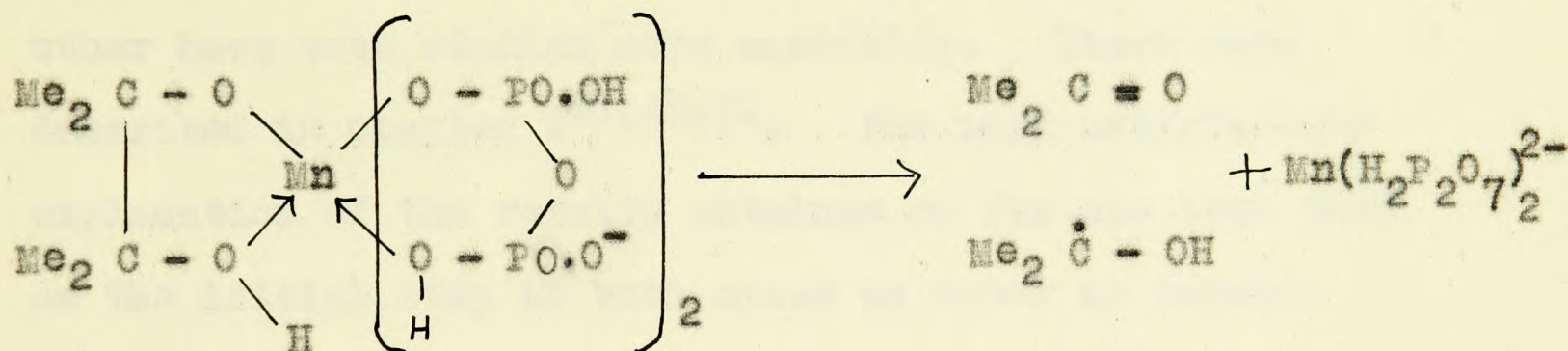
$$\frac{-d[\text{Mn}^{\text{III}}]}{dt} = \frac{k [\text{Hydroxy acid}] [\text{Mn}^{\text{III}}]}{a + [\text{Hydroxy acid}]}$$

at constant pH. The rate of reaction was reduced by a factor approaching 0.5 in the presence of vinyl cyanide, which is polymerised. The rate was also retarded by manganous ions. Mercuric salts were reduced to the mercurous state in situ; oxygen had no effect on the reaction rate. These observations indicate a mechanism involving free radicals, as is to be expected when a one-electron oxidising agent is used.

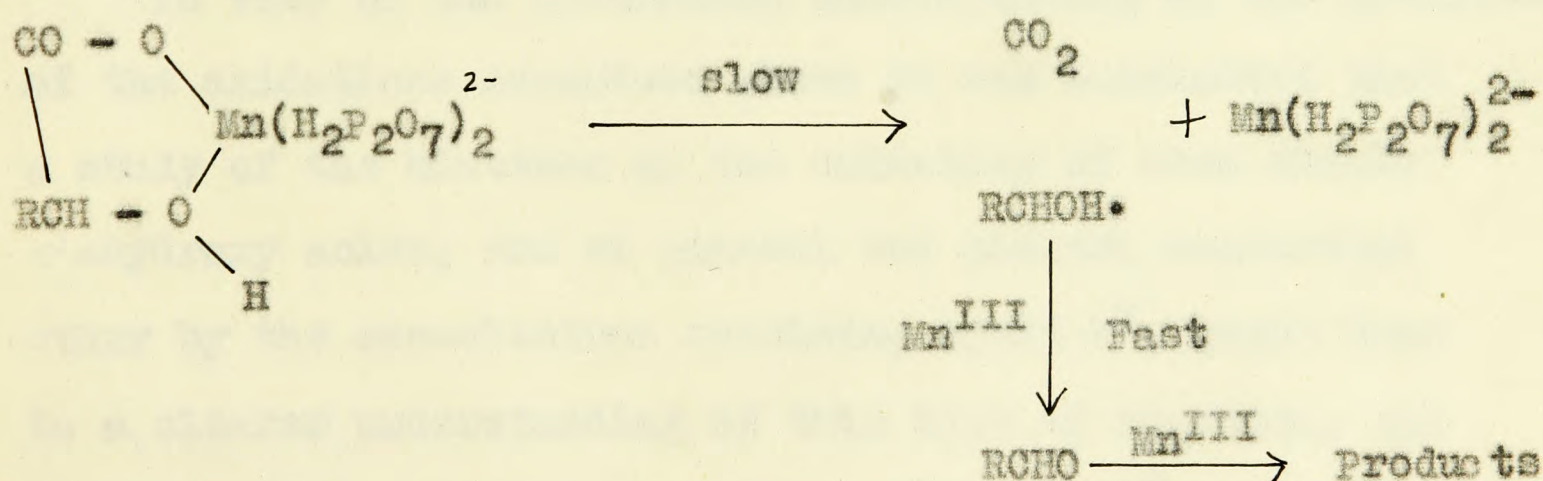
The kinetics were similar to those found by Drummond and Waters⁵³ for the oxidation of pinacol to acetone by

manganic pyrophosphate, with the exception of the manganous ion effect which was not observed in the pinacol oxidation. The reaction is catalysed by acid but the effect is somewhat complex; a graph of rate constant plotted against hydrogen ion concentration is sigmoid in form. The radicals formed are most likely $\cdot\text{CMe}_2\text{OH}$, since oxygen radicals such as $\text{Me}_2\text{C} - \text{C}(\text{OH})\text{Me}_2$ would probably not reduce mercuric salts.

The following reaction scheme was proposed:-



A similar scheme for the oxidation of α -hydroxy acids⁵⁵ is:-



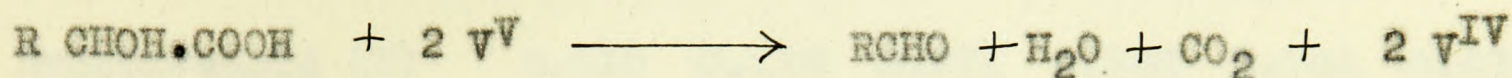
No satisfactory explanation of the manganous ion effect was proposed.

Some work has also been done in investigating the oxidation of lactic acid by chromic acid⁷⁶. This work only showed that the main products are acetaldehyde and carbon dioxide and that the kinetics of the reaction are complex; manganous ions were found to increase the rate of reaction somewhat.

The Cr^{VI} oxidations of pinacol and of its monomethyl ether have been studied more carefully. These were described in Chapter 2^{47,48,51}. The most satisfactory explanation of the results obtained so far has been that as the initial step in both cases an ester is formed between the substrate and chromic acid, but not all authors are yet in agreement on this point.

In view of the incomplete understanding of the mechanism of the oxidations described above it was considered that a study of the kinetics of the oxidation of some simple α -hydroxy acids, and of pinacol and pinacol monomethyl ether by the one-electron oxidising agent V^{V} , might lead to a clearer understanding of this type of reaction, and might provide a useful comparison with the V^{V} oxidations of simple alcohols and glycols which were described in Chapter 1.

The three α -hydroxy acids lactic acid, malic acid and mandelic acid were found to be oxidised rapidly at 25°C according to the equation



The aldehyde products were thereafter oxidised more gradually. Both lactic acid and malic acid yielded acetaldehyde while mandelic acid gave benzaldehyde. The α -substituted hydroxy acid α -hydroxyisobutyric acid consumed only 2 equivalents of V^{V} and yielded acetone.

Kinetic measurements were carried out in mixtures of perchloric acid and sodium perchlorate of constant ionic strength and are summarised in the adjoining tables.

Tables 3 (a) - (e) show that the reaction is of first order with respect to the hydroxy acids at both 1.0 N and 3.0 N acid. Tables 4 (a) - (c) show that lactic, malic and mandelic acids have an acid catalysed oxidation which can be expressed as

$$\text{Rate} = k + ah_0$$

where h_0 is the appropriate value of the Hammett acidity function. This type of acidity dependence was suggested for the V^{V} oxidation of pinacol by Littler and Waters⁴ who did not keep the ionic strength of their solutions constant

TABLE 3.

Oxidation of α -hydroxy-acids by vanadium (V): reaction order with respect to the α -hydroxy-acid at 25°

(a) Lactic acid in 1.0 N-perchloric acid; ionic strength,
S = 3.1 M

Hydroxy-acid (N)	0.110	0.219	0.274	0.329	0.384	0.439
k (10^{-4}sec.^{-1})	1.18	2.62	3.27	4.03	4.73	5.25
$10^4k/[\text{Acid}]$	10.8	11.1	12.2	12.3	12.3	12.0

(b) Lactic acid in 3.0 N-perchloric acid; ionic strength,
S = 3.1 M

Hydroxy-acid (N)	0.086	0.161	0.240	0.320	0.400
k (10^{-4}sec.^{-1})	3.77	7.60	11.1	15.8	19.0
$10^4k/[\text{Acid}]$	47.1	47.5	46.3	49.4	47.5

(c) Malic acid in 1.0 N-perchloric acid; no added salt.

Hydroxy-acid (N)	0.078	0.156	0.233	0.311	0.389
k (10^{-4}sec.^{-1})	1.05	2.17	3.30	4.53	5.60
$10^4k/[\text{Acid}]$	13.5	13.9	14.2	14.6	14.4

(d) Malic acid in 3.0 N-perchloric acid; no added salt.

Hydroxy-acid (N)	0.078	0.156	0.233	0.311	0.389
k (10^{-4}sec.^{-1})	4.22	10.0	14.4	19.5	24.0
$10^4k/[\text{Acid}]$	54.1	64.1	61.8	62.7	61.7

(e) α -Hydroxyisobutyric acid in 1.0 N-perchloric acid;
ionic strength, S = 3.1 M.

Hydroxy-acid (N)	0.096	0.121	0.145	0.193	0.241
k (10^{-4}sec.^{-1})	0.628	0.753	0.925	1.27	1.57
$10^4k/[\text{Acid}]$	6.51	6.25	6.40	6.58	6.51

TABLE 4

Dependence of oxidation rate on acidity;

ionic strength S = 5.1 M; Temp. = 25°

(a) Lactic acid (0.186M)

HClO ₄ (N) ⁴	h ₀	k(10 ⁻⁴ sec. ⁻¹)	
		Found	Calc.
0.5	12.9	2.42	2.37
1.0	28.2	3.30	3.77
1.5	47.9	5.8	5.56
2.0	69.2	7.3	7.50
2.5	100.0	10.5	10.3
3.0	138	13.8	13.8
4.0	282	25.7	26.9

Calc. k = 1.2 + 0.091h₀

(b) Malic acid (0.194M)

HClO ₄ (N) ⁴	h ₀	k(10 ⁻⁴ sec. ⁻¹)	
		Found	Calc.
0.5	12.9	3.17	3.11
1.0	28.2	4.58	4.77
1.8	60.3	8.05	8.3
2.6	105	13.5	13.1
3.4	186	21.2	21.0
3.8	246	28.3	28.5
4.6	427	51.7	43.2

Calc. k = 1.7 + 0.109h₀

(c) Mandelic acid (0.151M)

HClO ₄ (N) ⁴	h ₀	k(10 ⁻⁴ sec. ⁻¹)	
		Found	Calc.
0.5	12.9	57.0	56.9
0.7	19.0	60.8	60.6
1.0	28.2	67.5	66.2
1.4	42.7	75.5	75.1
1.8	60.3	83.3	85.8
2.2	79.4	98.2	97.5
2.6	105	109	113

Calc. k = 49.0 + 0.611h₀

(d) Pinacol (0.125M)

HClO ₄ (N) ⁴	h ₀	k(10 ⁻⁴ sec. ⁻¹)	
		Found	Calc.
0.5	12.9	43.5	43.3
1.0	28.2	51.7	48.1
1.8	60.3	55.7	56.2
2.6	105	70.8	67.6
3.4	186	87.8	83.1
4.2	324	112	123
(5.0	589	165 ± 10	191)

Calc. k = 41.0 + 0.253h₀

and used sulphuric acid which itself forms complexes with V^V . Table 4(d) shows that pinacol behaves similarly to the hydroxy acids in its reaction with V^V and confirms the views of Littler and Waters.

The rate constants given were all obtained from a first order plot of the disappearance of V^V . Straight lines were obtained for at least 85% of the reaction in all cases.

The overall kinetic equation is therefore

$$-\frac{d[V^V]}{dt} = k [\text{Substrate}] [V^V] (1 + ah_0)$$

In contrast to this, the kinetics of the V^V oxidation of pinacol monomethyl ether have been found to follow the rate equation

$$-\frac{d[V^V]}{dt} = k [\text{Ether}] [V^V] [H^+]$$

where k has a much smaller value than that for the hydroxy acid oxidations. The dependence of the rate of oxidation on the acid concentration is shown in Table 5 (These rates were kindly measured by J.S.Littler who used a "Unicam" ultra violet spectrophotometer to follow the increase in V^{IV} concentration²⁷)

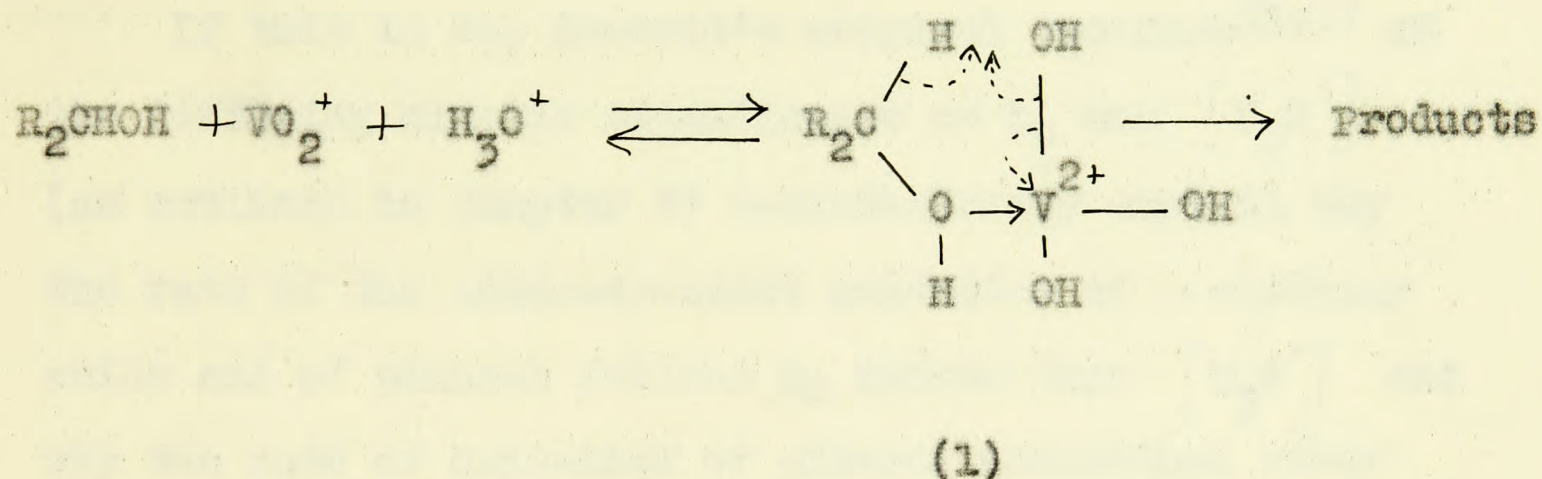
TABLE 5.Oxidation of pinacol monomethyl ether (0.111M) at 50°.

HClO ₄ (N)	1.0	1.8	2.6	3.4	4.2	5.0
k (10 ⁻⁴ sec. ⁻¹)	(1.52)	3.53	4.8	6.31	7.75	9.55
10 ⁴ k/ [H ⁺]	(1.52)	1.96	1.85	1.86	1.85	1.91
Weighted mean value: k = (1.63 ± 0.02) × 10 ⁻³ . [H ₃ O ⁺][Ether].						

None of these kinetic measurements give any evidence for initial complex formation between V^V and substrate such as is found in oxidations of hydroxy acids etc. by manganic pyrophosphate. This is not surprising, since the oxidation of cyclohexanol by V^V, in which complex formation is involved, shows no deviation from first order kinetics with respect to the alcohol⁴. Another difference between the V^V and Mn^{III} oxidations of hydroxy acids is that the manganous ion was found to have no significant effect on the rate of the V^V oxidation. It is possible that the manganic pyrophosphate solution contained a certain amount of Mn^{IV} formed by disproportionation of Mn^{III} into Mn^{II} and Mn^{IV}, and that the effect of adding manganous ions was to reduce the amount of Mn^{IV} and hence partly to eliminate oxidation by this species.

The V^V oxidations of cyclohexanol and of those secondary

α -glycols which are attacked at C - H bonds⁵ have been found to be linearly dependent on hydrogen ion concentration, and not on the Hammett function h_0 . For the alcohol oxidation it has been suggested⁴ that the reactive complex has a structure (1) in which both a monohydric alcohol molecule and a hydrated proton are co-ordinated to the VO_2^+ cation.

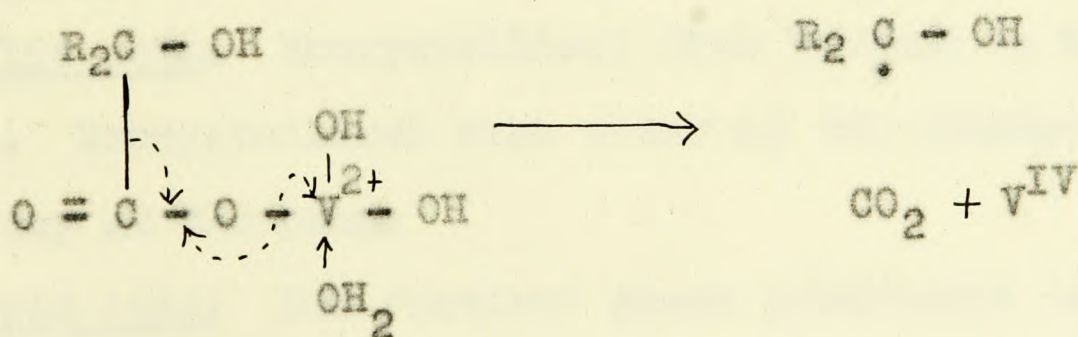


However in the oxidation of both pinacol and the α -hydroxy acids the complexes (II) and (III) which can undergo C - C fission, need not contain an extra water molecule, for the protonated structure (III) postulated for the acid catalysed reaction can be built from VO_2^+ , a substrate molecule, and a proton.

that the reacting species in the acid-dependent reaction is a doubly charged cation III.

It seems likely then that the reversible formation of complexes II and IV precedes the rate-determining, concerted electron redistribution which, in converting vanadium (V) into vanadium (IV), effects C - C fission giving radicals $R_2C(OH)^\bullet$ or $R_2C(OMe)^\bullet$, which are then rapidly oxidised further.

The acidity dependence of the α -hydroxyisobutyric acid oxidation, however, is intermediate between these two cases as is seen from Graph 1.; although it tends towards h_0 dependence at high acidities there is a much closer dependence on $[H_3O^+]$ at lower acidities than with the other hydroxy acids. It is possible that a reaction of the following type can occur in addition to the normal cyclic complex decomposition.



This would have a linear dependence on hydrogen ion concentration using the Hammett theory. This explanation cannot easily be reconciled with the comparative slowness of the reaction and the case is best left as an anomaly at

present.

The theory of formation of soluble complexes between the α -hydroxy acid and V^V is supported by the observation that, though the gradual acidification of aqueous ammonium metavanadate with solutions of lactic, malic, acetic and monochloroacetic acids leads in all cases to the separation of a brown precipitate (V_2O_5 polymer), the latter is easily soluble in an excess of each of the hydroxy acids but is not dissolved by further aqueous acetic or monochloroacetic acids, though the latter has a pK_a similar to those of the hydroxy acids.

EXPERIMENTAL

Materials

Lactic Acid. BDH "AnalaR" grade.

± Malic Acid. Recrystallised from ethyl acetate to Mpt 128° .

± Mandelic Acid. Recrystallised from benzene to Mpt 118° .

Pinacol. Recrystallised from water as the hexahydrate and sucked dry at the pump.

Perchloric Acid. BDH "AnalaR" grade perchloric acid was used to make up stock solutions of V^V . For other purposes Harrington's 60% perchloric acid was used.

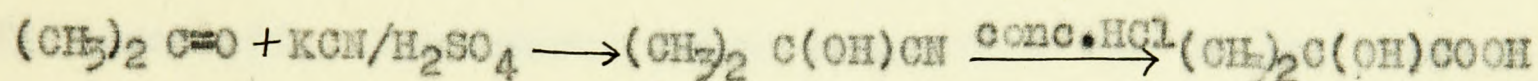
Sodium Perchlorate. Hopkins and Williams "recrystallised" grade.

Ammonium Metavanadate. B.D.H. "AnalaR" grade. (98% pure)

Ferrous Ammonium Sulphate. B.D.H. "AnalaR" grade.

Nitrogen was freed from any traces of oxygen by bubbling twice through wash bottles containing acid solutions of chromous sulphate in contact with zinc amalgam, and then through concentrated sulphuric acid.

Preparation of α -Hydroxyisobutyric acid.



(a) Preparation of Acetone Cyanhydrin⁷⁸

B.D.H. "AnalaR" grade acetone (58g. 1 mole) was added to potassium cyanide (62g. 1 mole) in 200 cc. water in a 1 l. 3-necked flask. 30% Sulphuric acid (333g.) was added dropwise during 1 hr. with constant mechanical stirring. The temperature was kept below 20° by adding ice from time to time. (This addition of ice also keeps the solution thin enough to stir, since potassium hydrogen sulphate is precipitated during the reaction). The solution was extracted with ether (3 x 100 c.c.), and the ether extract was dried over magnesium sulphate. After evaporation of ether, the product was distilled and the fraction of B.Pt. 82-83°/18mm. was collected. Yield 43.5g. (51%), but more

material was obtained by further ether extraction.

(b) Conversion of Acetone Cyanhydrin to α -Hydroxyisobutyric Acid.⁷⁹

Acetone Cyanhydrin (43.5g.) was left for three days with 250 c.c. concentrated hydrochloric acid. Some of the hydrogen chloride was then evaporated and after cooling the solution, some ammonium chloride which had separated was filtered off. Continuous ether extraction for 24 hr. removed 39.0g. of α -hydroxyisobutyric acid (73% yield). The product was recrystallised from petroleum ether (B.Pt. 60°-80°).

M.Pt. 79°. Found: C, 46.27; H, 7.75.

Calc. for $C_4H_8O_3$: C, 46.15; H, 7.75%.

Hammett Acidity Function

The values of h_0 cited in table 2 were obtained from Paul and Long's review³⁴ and refer to mixtures of perchloric acid and sodium perchlorate of ionic strength $S = 6$.

Values for $S = 5.1$ are not available but should be directly proportional to those given, except possibly in 5.0 M perchloric acid to which no sodium perchlorate had been added.

Standard solutions.

Solutions of the hydroxy acids and of perchloric acid

were standardised by titration against N. sodium hydroxide solution, which was itself standardised by titration against standard potassium hydrogen phthalate solution.

Sodium perchlorate was used as an 8M solution in distilled water. 10 C.c. portions of this solution were carefully evaporated in a crucible covered by a watch glass and dried to constant weight in order to check the strength of the solution. An 8M solution was found to have a density of 1.605g./c.c.

The vanadate solution for kinetic oxidation work was prepared by dissolving 6.0g. ammonium metavanadate in 100 c.c. boiling distilled water. The hot solution was slowly poured into the calculated quantity of cold perchloric acid with stirring to prevent precipitation of vanadium pentoxide. Care was taken to prevent the temperature of the acid solution rising above 50° since this leads to decomposition of the solution and precipitation of vanadium pentoxide. The solution was made up to 250 c.c. to produce an 0.2M solution of V^V . Solutions of V^V for titration and uptake experiments were prepared in sulphuric acid and were found to have great stability. They did not deteriorate to any detectable extent in several months. Standardisation was effected by titration against ferrous ammonium sulphate solution which itself was standardised against potassium

dichromate solution made up from the pure solid.

Titration Procedure.

A typical reaction mixture had an initial volume of 100 c.c. made up with 25 c.c. of 0.2M V^V solution and measured volumes of organic substrate solution and sodium perchlorate solution which served to adjust the ionic strength to the required value. Additional perchloric acid and distilled water were added as necessary. 10 C.c. portions of the mixture were withdrawn in a pipette at suitable intervals of time and were added to 5 c.c. portions of 0.1 N ferrous ammonium sulphate to quench the reaction. The mixture was diluted to about 25 c.c. with distilled water and 5 c.c. of orthophosphoric acid was added together with 2 drops of barium diphenylamine sulphonate solution (0.3g. in 100 c.c. aqueous solution). This solution was then titrated against 0.1 N V^V solution in sulphuric acid from a microburette until a permanent blue colour appeared. Straight line graphs were obtained by plotting $\log_e$ (concentration of V^V) against time t . First order rate constants k_1 were calculated from the slope of these graphs.

$$\text{Rate} = \frac{-d[V^V]}{dt} = k_1 [V^V]$$

$$\therefore \frac{-d[V^V]}{[V^V]} = k_1 dt$$

$$\therefore -\log_e [V^V] = k_1 t + \text{Constant.}$$

Temperature Control.

Kinetic oxidations were performed in a 250 cc. conical flask, corked, and immersed in a thermostat maintained at $25^\circ \pm 0.1^\circ$ by means of a toluene-mercury regulator.

Reactions under Nitrogen

The nitrogen was first passed into a wash bottle containing all the components of the reaction mixture except V^V , and then into the reaction itself. This was contained in a 250 c.c. 3-neck round bottom "Quickfit" flask. The nitrogen escaped from the flask through a constricted piece of rubber tube to prevent air diffusing back. The third neck of the flask was used for removing samples. During periods when this neck had to be opened for such purposes, nitrogen was allowed to bubble more quickly through the reaction mixture.

Uptake Experiments

A standard solution of 0.2 M ammonium metavanadate in 1 M sulphuric acid (200 c.c.) was mixed with 5 or 10 c.c. of the standard hydroxy acid solution leaving at least 50% excess V^V than would be required for complete oxidation. The mixture was left in a thermostat at 25° . At intervals of 24 hr. or less 10 c.c. portions were removed and titrated as described above. The number of moles of V^V used per mole of organic substrate was calculated from the original and final strengths of the reaction mixture. In some cases the mixture was finally heated on a boiling water bath for a few hours to ensure completion of the reaction, but this procedure results in volatilisation of low boiling intermediates such as acetaldehyde, formaldehyde etc.

The maximum quantities of V^V absorbed per mole of hydroxy acid were: lactic acid 4.6 moles, malic acid 4.0 moles, mandelic acid 2.9 moles, α -hydroxyisobutyric acid 2.0 moles.

The Effect of Manganous Ions on Reaction Rates

Reaction mixtures of similar constitution to those described above were made up to include manganous sulphate

in sufficient amount to give 0.05 or 0.10 M solutions. These reactions were compared with reactions containing magnesium sulphate in similar quantities instead of manganous sulphate, the other components being identical. In all cases the reaction rates of similar solutions containing the two salts were found to be within 2% of each other.

Products of Reaction

The following procedure was found to be convenient for isolating solid derivatives of volatile aldehydic or ketonic products of oxidation. A solution of ammonium metavanadate in sulphuric acid was stirred in a flask which was heated on a water bath. Vapours from above the solution were sucked out through a neck of the flask and bubbled slowly through a wash bottle trap containing the necessary reagent solution such as dimedone or 2,4 dinitrophenylhydrazine. The organic substrate was then introduced into the V^V solution and a precipitate soon formed in the bubbler if a carbonyl product was formed. The derivative was recrystallised to constant melting point.

Both lactic acid and malic acid gave a dimedone derivative of M.Pt. 139° and mixed M.Pt. with the acetalde-

hyde derivative 139° .

Mandelic acid gave benzaldehyde which had a characteristic smell and a 2,4 dinitrophenylhydrazone of M.Pt. 235° which was not depressed by admixture with the derivative of authentic benzaldehyde.

CHAPTER 5

OXIDATION OF OXALIC ACID.

Oxalic acid is one of the simplest of organic compounds, but its oxidation by various reagents has many complexities.

The permanganate oxidation of oxalic acid⁵⁶ which is catalysed by manganous ions and retarded by excess sodium oxalate, was described in Chapter 2.

Dhar⁸⁰ investigated the chromic acid oxidation of oxalic acid and found that although the addition of sulphuric acid increased the rate of reaction at first, above about 0.5 M sulphuric acid the effect of further addition of acid was to decrease the rate of reaction.

Bobtelsky and Glasner³ have investigated the oxidation of oxalic acid by vanadium (V) in various mineral acids and have reported that the reaction has a negative acid catalysis in the range 0-3.5 M acid. At about 3.5 M acid the rate of oxidation passes through a minimum and the acid catalysis becomes positive at greater acidities than this. The oxidation was of first order with respect to V^V concentration, but was of approximately second order with respect to oxalic acid concentration in solutions of various concentrations of sulphuric acid. It was found that manganous ions accelerated the reaction, the half life of the reaction being roughly

proportional to $\frac{1}{[\text{Mn}^{2+}]}$. This effect was also found by Dhar in the Cr^{VI} oxidation of oxalic acid. Bobtelsky and Glasner also found that ferric salts decreased the rate of the reaction. In a darkened flask the half life was roughly proportional to $[\text{Fe}^{3+}]$, but in sunlight the reaction was accelerated in the presence of ferric ions. This photochemical decomposition of ferric oxalate to give ferrous ions (which rapidly reduce oxidising agents) is well known and has been studied in some detail by Horne³².

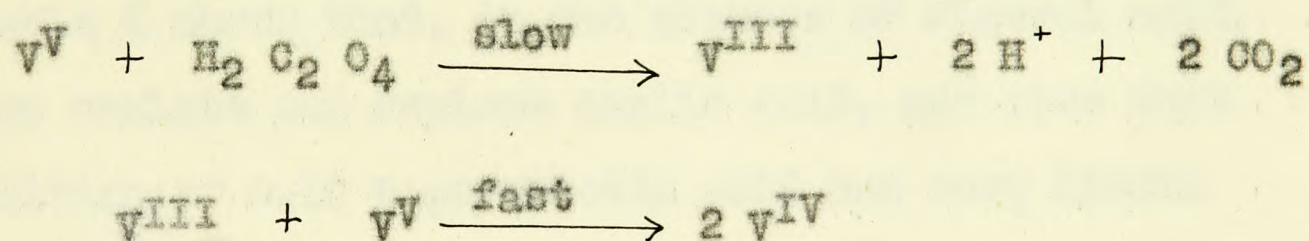
Since both the negative acid catalysis and the reaction order with respect to oxalic acid appeared to be unique among oxidations by V^{V} it was decided to investigate the V^{V} oxidation of oxalic acid in more detail. The results, which accord with the account given by Bobtelsky and Glasner, show that the unusual features of the reaction are due to the formation of a series of complexes between vanadium (V) and oxalic acid, the stabilities of which are pH dependent.

Experimental Results

The stoichiometry of the overall reaction accords with the equation



This has also been reported by West and Skoog¹² and by Morette and Gaudefroy³¹. It was not possible to adduce any experimental evidence of the transient formation of organic radicals or radical-ions such as $\text{H C}_2\text{O}_4^\bullet$, $^\bullet\text{C}_2\text{O}_4^-$, or $^\bullet\text{CO}_2^-$, since this oxidation of oxalic acid, unlike those effected by manganic salts¹⁵, or by Fenton's reagent³³, does not promote the chain reaction between oxalic acid and mercuric chloride or the polymerisation of vinyl cyanide. However this lack of evidence of radical formation does not necessarily indicate that a 2-electron oxidation, such as



is involved; the oxalic acid radical may be oxidised so rapidly by V^{V} that its independent existence cannot be demonstrated by the tests mentioned above.

The oxidations of oxalic acid by vanadium (V) were all carried out in solutions in which oxalic acid was present in large excess over oxidant, and were all found to follow strict first order kinetics with respect to the vanadium for at least 85% of the reaction.

The general trend of the rate of oxidation in the presence of perchloric acid is shown in Graph 2. Graph 3

shows the relationship between the rate of oxidation and the initial concentration of the oxalic acid in solutions made up from ammonium metavanadate with no added mineral acid, whilst the curves of Graph 4 show corresponding rates of oxidation in the presence of 0.5, 1.0 and 1.45 M perchloric acid respectively. Graph 5 shows that the results plotted in Graph 4 correspond approximately to a reaction having second order kinetics with respect to total oxalic acid concentration, especially at the lower values and serve to explain how Bobtelsky and Glasner came to their conclusions mentioned above.

Table 6 shows that, in the absence of mineral acid, ammonium oxalate can replace oxalic acid, and also that the addition of 0.10 M perchloric acid has very little effect on the oxidation rate. Table 7, however, shows how slightly larger amounts of perchloric acid decrease the oxidation rate when an excess of oxalic acid is present.

It is clear from Graph 3 that in the pH range 1 - 2 a complex of vanadium (V) with two molecules of oxalic acid must be reacting with a third molecule of oxalic acid before any vanadium (IV) is formed, but the potentiometric titration of oxalic acid with alkali in the presence of ammonium metavanadate (Graph 6) shows that the complex formation effects no significant pH change in this region. Graph 7 shows a similar titration using only half the amount

of vanadium used to obtain Graph 6. It is apparent that sections (3) and (4) of the curves of these graphs are dependent on the presence of ammonium metavanadate. An interpretation of these curves is only possible if it is assumed that a complex containing one molecule of vanadium and two molecules of oxalic acid is present initially. The use of sodium orthovanadate instead of the ammonium salt resulted in a decrease in the length of section (4) of the curves from 2 to 1 equivalents of alkali per equivalent of V^V .

Graph 4 indicates that the vanadium (V)-oxalic acid complex, formed in neutral solution, breaks down gradually to other complexes as the acidity of the medium is increased, and finally Table 8 indicates that at high acidity (4.2 M) the reaction approaches one of first order with respect to oxalic acid even at low concentrations of the latter. As Graph 2 shows, the oxidation then becomes an acid catalysed process, just as would be expected from a reaction involving $V(OH)_3^{2+}$ rather than VO_2^+ .

Table 6.

Oxidation of Oxalic Acid by Aqueous Ammonium Metavanadate (0.05 M) at 40°.

Ionic strength of all solutions = 2.1.

k_1 = First order rate constants, (sec^{-1})

<u>[Oxalic Acid]</u> (M)	only NaClO_4 added.	$(\text{NH}_4)_2\text{C}_2\text{O}_4$ added (0.05M)	HClO_4 added (0.10M)
0.101	$0.125 \cdot 10^{-3}$	$0.320 \cdot 10^{-3}$	$0.22 \cdot 10^{-3}$
0.201	$1.18 \cdot 10^{-3}$	$1.77 \cdot 10^{-3}$	$1.08 \cdot 10^{-3}$
0.401	$4.15 \cdot 10^{-3}$	—	$4.17 \cdot 10^{-3}$

Table 7.

Effect of varying Perchloric Acid concentrations in the range 0.0 - 1.0 M.

($[\text{V}^{\text{V}}] = 0.05 \text{ M.}$, $[\text{H}_2\text{C}_2\text{O}_4] = 0.121 \text{ M.}$, Ionic Strength = 2.1, Temp. = 40°C)

$[\text{HClO}_4]$ (M)	$10^4 \times k_1$ (sec^{-1})	$[\text{HClO}_4]$ (M)	$10^4 \times k_1$ (sec^{-1})
0.00	3.17	0.40	1.98
0.10	2.92	0.50	1.83
0.20	2.50	0.80	1.60
0.30	2.22	1.00	1.44

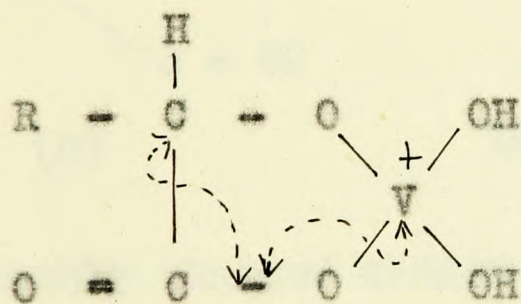
Table 8. Reaction order with respect to Oxalic Acid in
4.2 M Perchloric Acid at 50°.

$$[V^V] = 0.05 \text{ M.}, \quad \text{Ionic strength} = 5.1$$

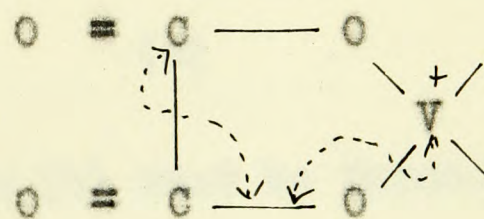
[Oxalic Acid], (M)	0.024	0.049	0.073	0.097	0.122
$10^3 \times k_2 = k_1/[H_2C_2O_4]$	2.17	2.29	2.42	2.60	2.71

Discussion

It was shown in Chapter 4 that the formation of a cyclic transition complex (e.g. I) favours the movement of an electron to a positively charged vanadium centre during the oxidation of hydroxy acids and certain glycols by vanadium (V)



I



II

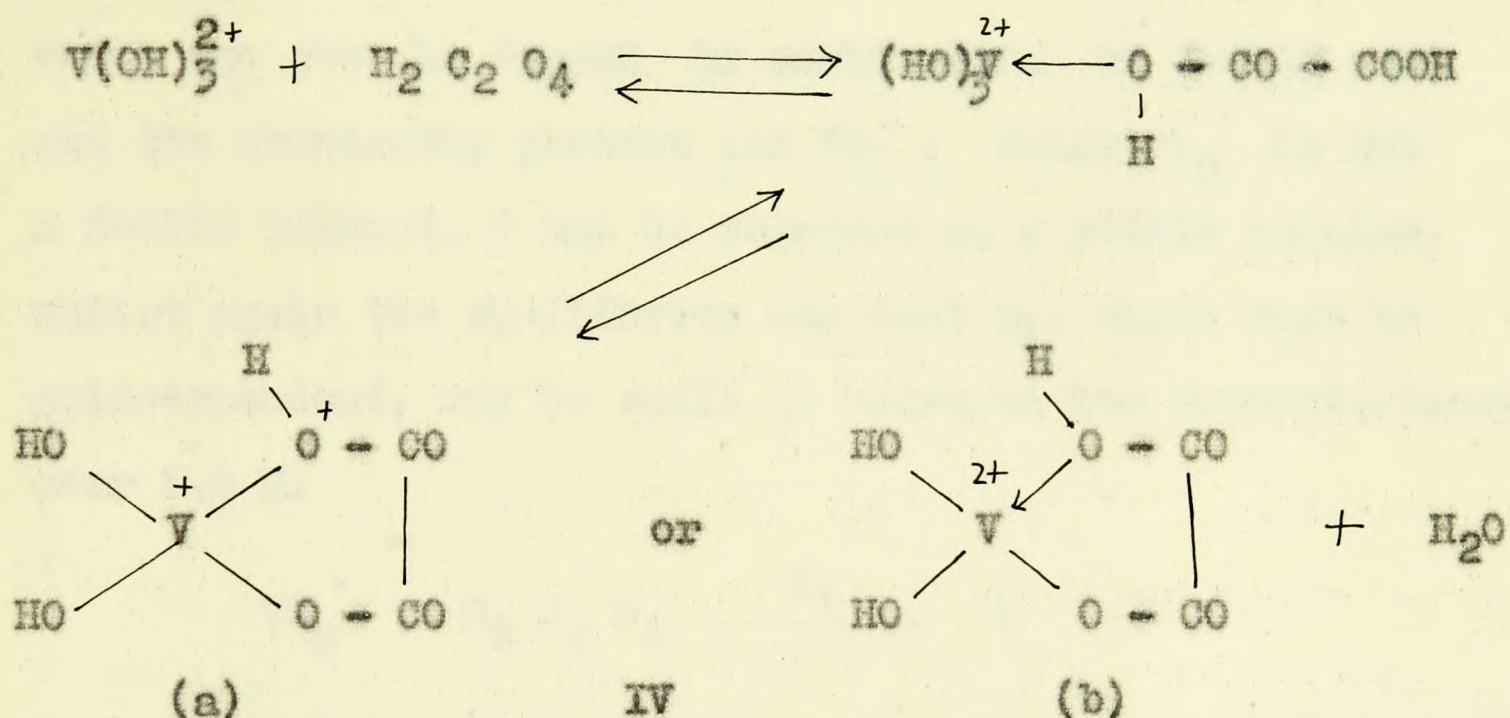
Consequently for oxalic acid oxidation a complex which can undergo electron transfer reactions is more likely to contain the cyclic group II rather than a linear group



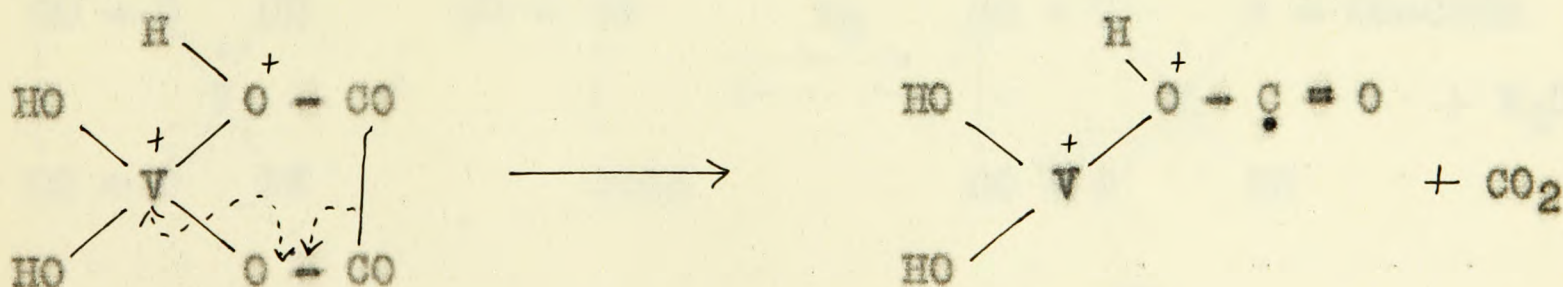
in which there is no structural assistance to 1- electron

movement.

Now at acidities above 4 N the oxidation of oxalic acid appears to be a reaction of normal type, i.e. of first order with respect to both $[H^+]$ and $[oxalic\ acid]$. Under these conditions it is rational to consider that the active oxidant is the uncomplexed ion $V(OH)_5^{2+}$ and to write the equilibria

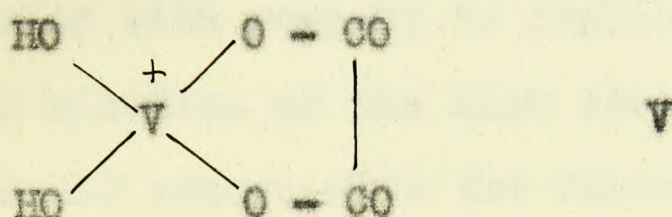


The doubly charged cation IV, (a) or (b) must be formed to a slight extent only and will undergo the 1-electron redistribution.

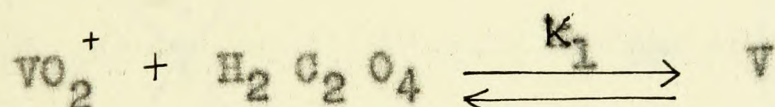


to a compound of vanadium (IV) having a bound radical,
 $\bullet\text{CO-OH}$ which would very easily reduce another ion of V^{V} .

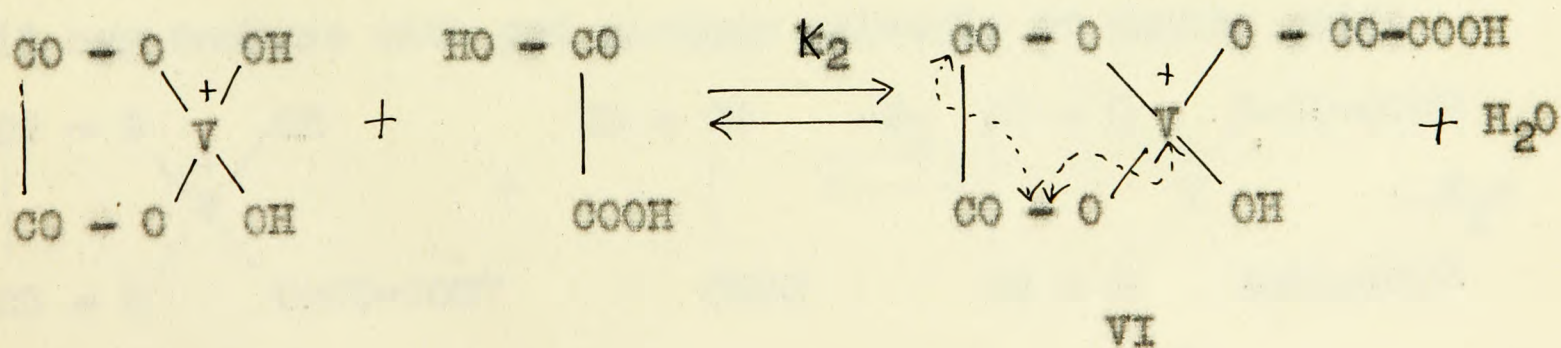
As the acidity is lowered the cation IV, being a
 strong acid would lose a proton giving a complex V



which can also be formed by combination of oxalic acid
 and the abundantly present ion VO_2^+ . Since VO_2^+ is but
 a feeble oxidant, V can be regarded as a stable complex,
 whilst again the equilibrium constant K_1 , which must be
 acid-dependent, may be small at hydrogen ion concentrations
 over 1.0 N.



However at lower acidities the hydroxyl groups of V may
 react with a further molecule of oxalic acid, and, since
 the $-\text{COOH}$ group is electrophilic, VI would be less stable
 than V and may decompose as shown.

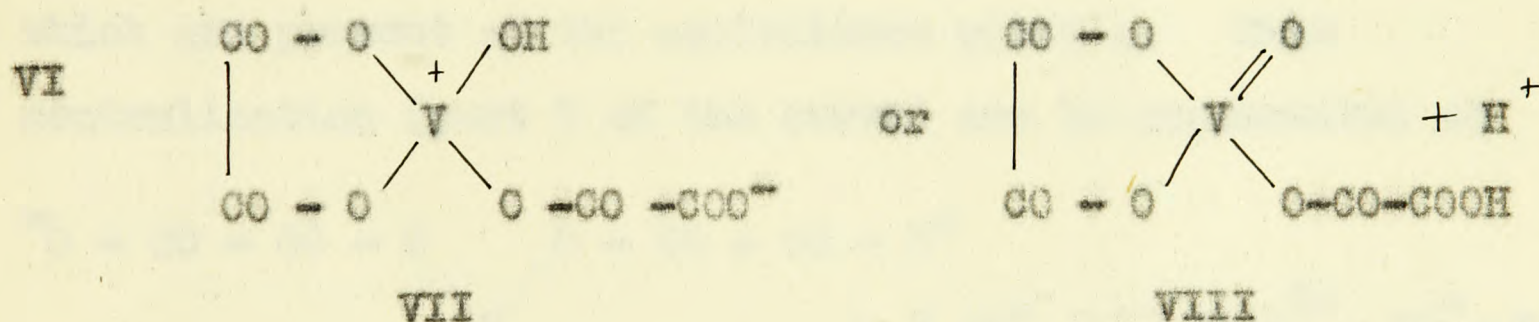


Under conditions in which both k_1 and k_2 are small

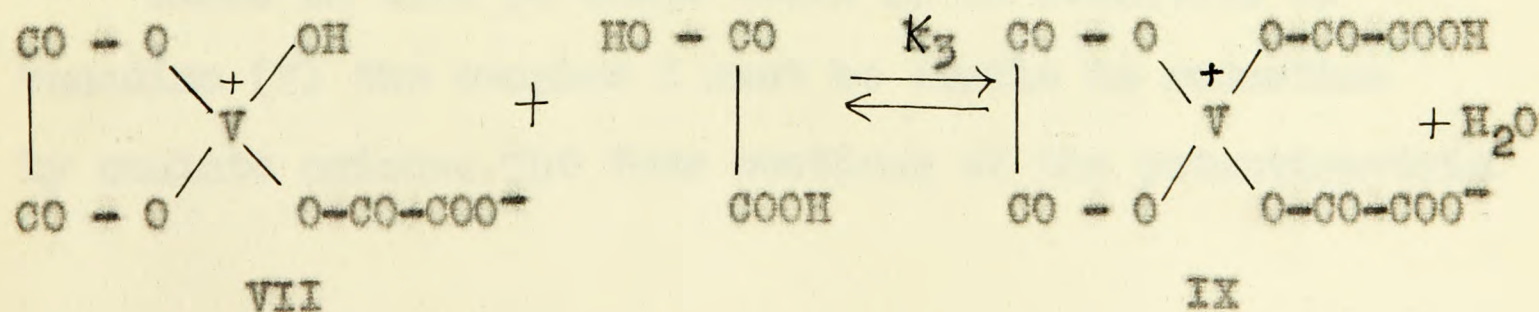
$$[V] = \text{approx. } k_1 [VO_2^+][H_2C_2O_4], \text{ and } [VI] = \text{approx. } k_1 k_2 [VO_2^+][H_2C_2O_4]^2$$

so that the decomposition of VI could appear to be roughly of second order with respect to oxalic acid. More exactly, however, the kinetics of the slow decomposition of the complex VI would accord with the experimental data shown in Graph 4.

As the acidity drops still further oxalic acid would begin to ionise, and likewise would VI; the latter would then give an electrically neutral complex VII or VIII in which the electron attracting properties of the vanadium atom would be very small.

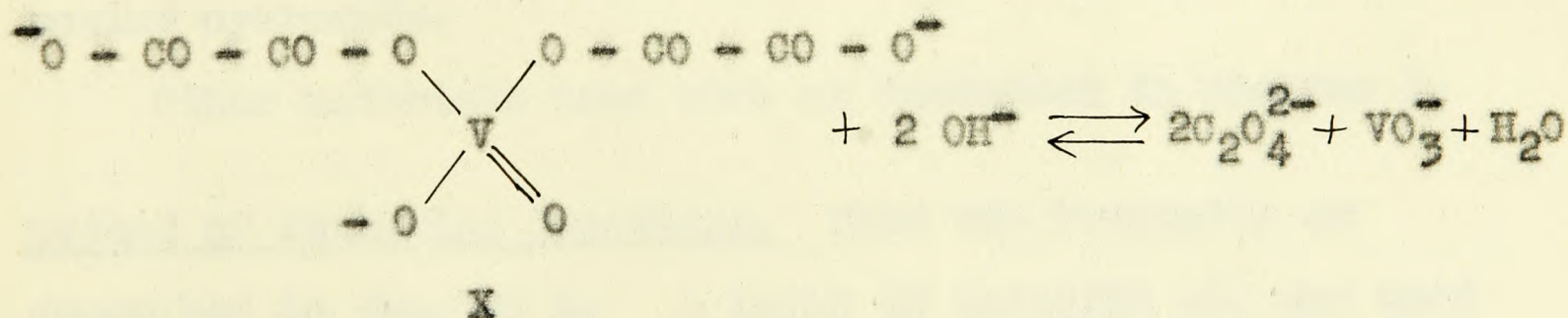


The tendency of this complex to lose oxalic acid would also diminish as the environment becomes more neutral. However it may combine with yet another molecule of oxalic acid.



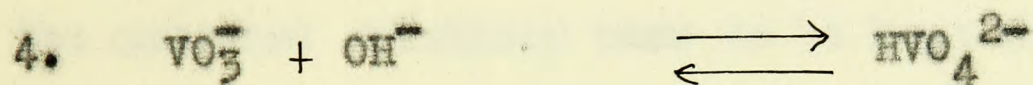
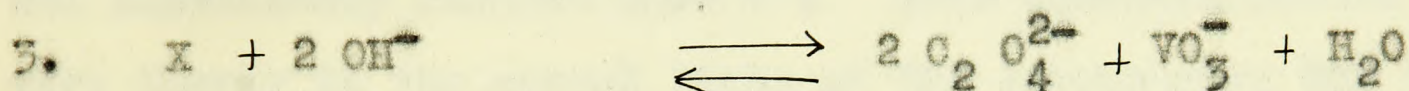
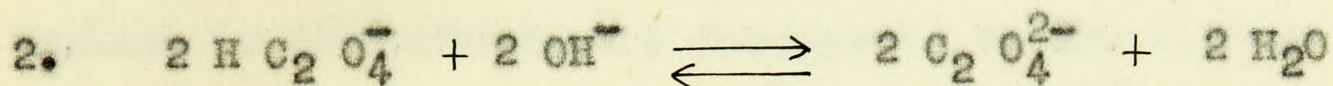
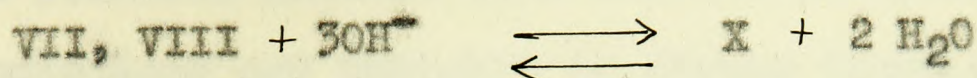
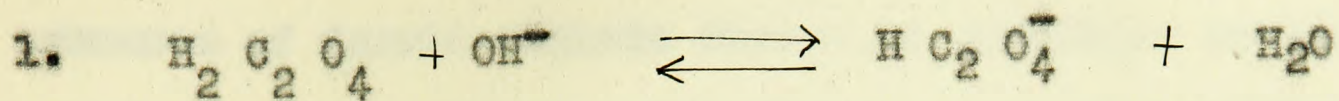
Graph 3 shows that a complex of formula $(\text{H}_2\text{C}_2\text{O}_4)_2 \cdot \text{V}^{\text{V}}$ is an oxidiser of oxalic acid so that it may be concluded that a cyclic tertiary complex such as IX easily decomposes to a compound of vanadium (IV). This is rational since IX, unlike VIII, necessarily has a positively charged vanadium atom at its centre.

Oxidation of oxalic acid by V^{V} does not occur at pH values above about 2. The potentiometric titrations (Graphs 6 and 7) indicate that the dissociation constant of VII is of the same order as that of oxalic acid, but that in the pH region 2 - 6 another complex (e.g. X) is formed which consumes two equivalents of alkali in decomposing to free oxalate, $\text{C}_2\text{O}_4^{2-}$ and metavanadate, VO_3^- , anions which are present at the equivalence point A. This neutralisation (part 3 of the curve) can be represented as



Since in this pH range there is no reduction of vanadium (V) the complex X must be stable to reduction by oxalate anions. The four sections of the potentiometric

titration curve in Graph 6 thus represent.



Experimental

Oxalic Acid: A weighed amount of B.D.H "AnalaR" grade material was dissolved in distilled water. The strength of the solution was checked by titration against standard sodium hydroxide.

Other materials used were as described in Chapter 4.

Method of Following Reactions. This was basically as described in Chapter 4. A layer of paraffin oil was used on top of the water in the thermostat in order to reduce loss by evaporation at the higher temperatures required by the reaction (40° or 50°), so that a reliable constant head device was unnecessary.

The analytical procedure involved titrating V^V in the presence of ferric oxalate formed in quenching portions of the reaction mixture with ferrous ammonium sulphate. Such solutions are unstable in strong light as the ferric oxalate is photochemically decomposed to the ferrous state, and immediately reduces more V^V . This reaction occurs very slowly in the normal light of the laboratory, but if the quenched solutions were to be kept for more than a few seconds before being titrated, they were stored in a dark cupboard, where they were found to be stable for several hours. The rates of the direct oxidations of oxalic acid by V^V were not affected by the exclusion of light. This was demonstrated by repeating some experiments in a blackened flask. Reaction also proceeded at the same rate if performed under nitrogen instead of in air.

When ammonium oxalate solution was added to some of the reaction solutions a white precipitate appeared under certain conditions, especially when high concentrations of sodium perchlorate were being used. The solid was filtered off and when heated on a spatula melted at first, charring slightly and then solidified to a white mass which gave a strong golden yellow flame. This white mass reacted alkaline to litmus and congo red papers and liberated a gas (CO_2) with dilute sulphuric acid; it was most likely sodium carbonate. The original solid crystals gave an aqueous

solution of pH about 3 and liberated carbon dioxide from sodium carbonate solution. Acid solutions of it were not coloured and did not therefore contain any V^V . It was concluded that the material was sodium hydrogen oxalate, which has a relatively low solubility in water. Because of this relative insolubility it was not possible to investigate the effect of added oxalate on the rate of reaction as thoroughly as was intended.

Potentiometric Titrations.

These were performed at room temperature, when the rate of oxidation of the oxalic acid is negligible. A Doran pH Meter was used with one glass electrode and one saturated calomel electrode in the titration vessel; 50 c.c. of solution was present initially and 1 N sodium hydroxide solution was run in from a burette in $\frac{1}{2}$ c.c. (or more) portions, the pH of the solution being measured after each addition of alkali.

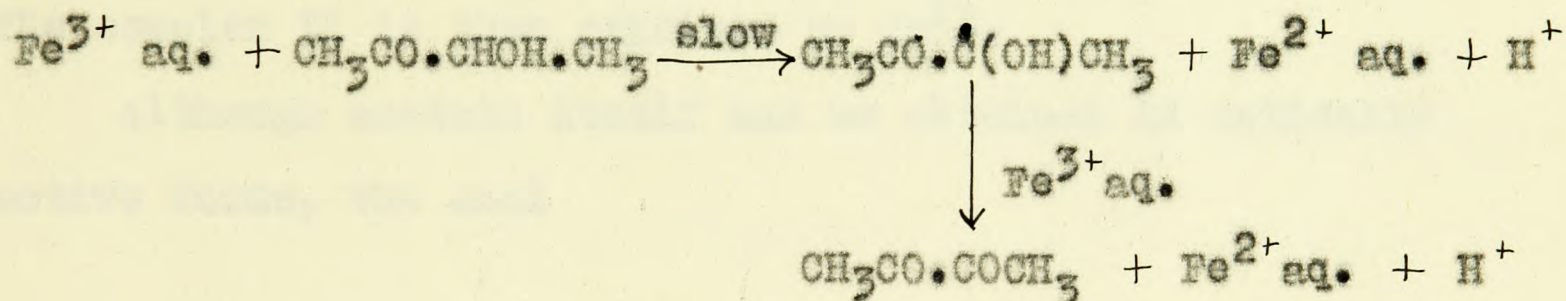
CHAPTER 6OXIDATION OF ACETOIN AND 3-METHYL 3-HYDROXY 2-BUTANONE(METHYL ACETOIN)

Acetoin is readily attacked by most oxidising agents but only in a few cases has its oxidation been investigated in any detail. These oxidations may be divided into two classes, (a) those which produce immediate C - C bond fission to give acetaldehyde and acetic acid, and (b) those that initially give diacetyl, which may, or may not, be oxidised further to acetic acid. Into the former class come the two-electron glycol splitting reagents lead tetraacetate and periodic acid. Lead tetraacetate has been reported⁸⁴ to give at least a 95% yield of acetaldehyde. Into class (b) come oxidations by air, iron (III), bismuth (III) and alkaline solutions of copper (II) complexes. The fact that acetoin readily absorbs oxygen⁸⁵ has to be remembered in any kinetic studies of its reactions. Rigby⁸⁶ isolated diacetyl in high yield from the bismuth (III) oxidation of acetoin in glacial acetic acid.

Thomas, Trudel and Bywater⁶⁶ obtained the following rate equation for the Fe^{III} oxidation of acetoin in perchloric acid solutions.

$$-\frac{d[\text{Fe}^{\text{III}}]}{dt} = \frac{k[\text{Fe}^{\text{III}}][\text{acetoin}]^{1.1}}{[\text{H}^+]}$$

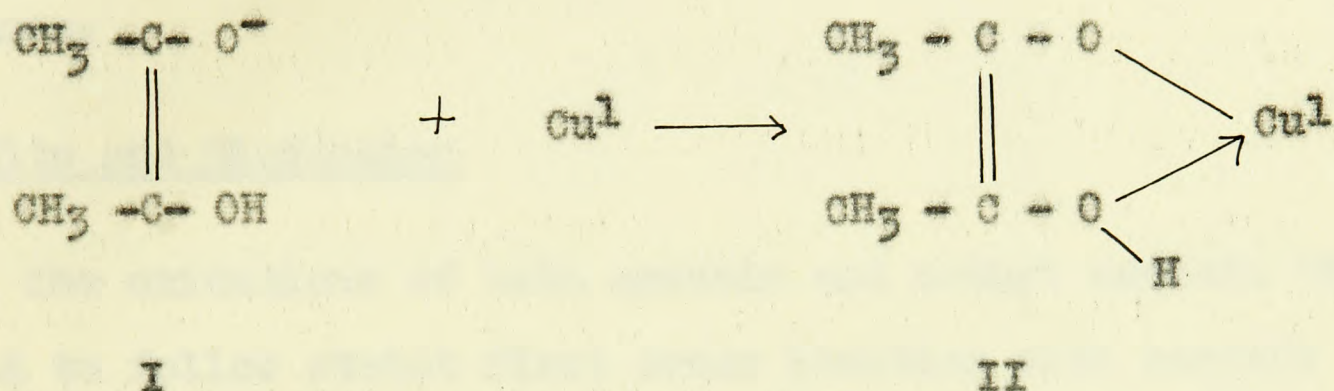
They also found that the reaction initiated the polymerisation of methyl methacrylate and that the rate was reduced in the presence of oxygen; these observations indicate the presence of free radicals. It was proposed that the ferric ion broke the C - H bond of the :CHOH group as the first stage of the reaction.



The acid dependence was taken to show that the species $\text{FeOH}^{2+} \text{ aq.}$ reacts faster than $\text{Fe}^{3+} \text{ aq.}$. The methyl substituted acetoin $(\text{CH}_3)_2 \text{COH} \cdot \text{CO} \cdot \text{CH}_3$ was not attacked by Fe^{III} . It is clear that this compound cannot take part in a reaction of type (b) above.

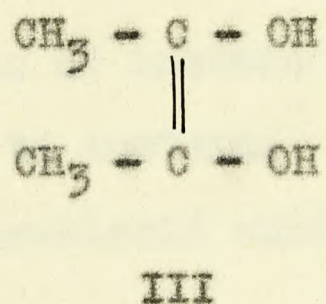
Marshall and Waters⁶⁷ have oxidised acetoin with complexed copper (II) salts in alkaline solution. This work was described in Chapter 2. They also proposed that the C - H bond of the :CHOH group was broken initially, but in alkaline solutions this is a fast base catalysed

reaction. They therefore proposed that the enolate anion I so formed, gave a chelated complex with copper (I) as the rate determining step.



The complex II is then oxidised by Cu^{II} .

Although acetoin itself can be obtained in optically active forms, the enol



derived from it, has no asymmetric centre and cannot possess optical activity. It has been observed⁸⁷ that solutions of optically active acetoin only lose their optical activity in the presence of alkali, this shows that the enol III is not formed in acid solutions.

Both acetoin and methyl acetoin are readily oxidised by

V^V in acid solutions. It was decided therefore to investigate these reactions in some detail and to see if there was any similarity between the kinetics of the oxidations and of those of either simple alcohols or simple ketones.

Results and Discussion

The oxidations of both acetoin and methyl acetoin were found to follow strict first order kinetics with respect to V^V for at least 85% of the reaction. Tables 9 and 10 show that the reaction is also of first order with respect to each of the hydroxy ketones. The acidity dependences of the rates of the two oxidations were dissimilar however.

Table 9. Oxidation of Acetoin by V^V. Reaction order with respect to Acetoin.

Reactions in 1.0M perchloric acid; ionic strength = 3.1M.

Temp. = 25°.

[Acetoin](M)	0.042	0.059	0.084	0.101	0.134	0.168
$10^3 k_1$ (sec ⁻¹)	1.25	1.68	2.58	3.00	4.22	5.83
$10^3 k_1 /$ [acetoin]	29.8	28.5	30.7	29.7	31.4	34.7

Table 10. Oxidation of 3-Hydroxy 3-methyl 2-butanone.
 Reaction order with respect to Hydroxy Ketone.
 Reaction in 1.0 M perchloric acid; ionic
 strength = 3.1 M, Temp. = 25°

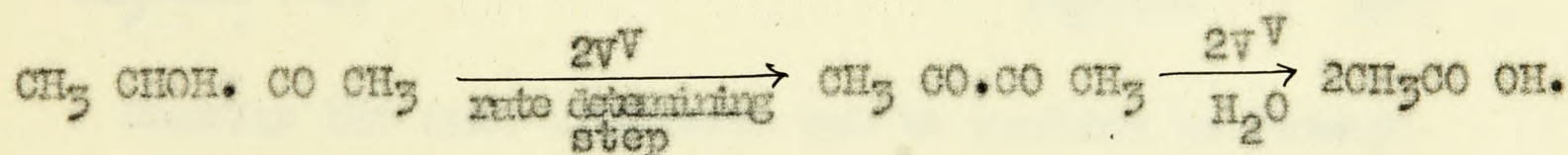
[Hydroxy Ketone] (M)	0.0405	0.054	0.081	0.135
$10^3 k_1$ (sec ⁻¹)	0.259	0.361	0.527	0.927
$10^3 k_1 / [\text{Hydroxy Ketone}]$	6.4	6.7	6.5	6.85

Graph 8(a) shows that the acetoin oxidation is entirely acid dependent. The rate of reaction is approximately proportional to the Hammett acidity function h_0 . Graph 8(b) is a plot of $\log_{10}(\text{rate})$ against H_0 ($= -\log_{10} h_0$) which should be a straight line of unit slope for a strict dependence on h_0 . It was not possible to use an ionic strength above 3.1 M for the reaction and two lines are given in graph 8 (b) using values of H_0 (taken from Paul and Long's Review³⁴) for mixtures of perchloric acid and sodium perchlorate of ionic strength 6 M, and for perchloric acid solutions without extra salt added, respectively. It will be seen that the former plot has a slope very nearly equal to unity.

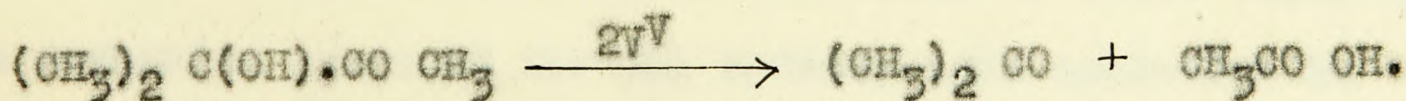
The acidity dependence of the methyl acetoin oxidation is shown in graph 9(a) (ionic strength = 5.1 M). The reaction has both an acid dependent component and an acid independent component. The acid dependent component is plotted

logarithmically against values of H_0 for 6M solutions of perchloric acid and sodium perchlorate in graph 9 (b) which is nearly a straight line of unit slope.

Acetoin rapidly consumes 4.0 equivalents of V^V in its oxidation. Partly oxidised reaction mixtures were noted to smell strongly of diacetyl, but this material is oxidised very much more rapidly than acetoin and it was not possible to isolate diacetyl from the reaction mixtures. Nor was it possible to obtain any evidence for other carbonyl compounds, such as acetaldehyde in the reaction products. Acetaldehyde is oxidised by V^V at a much slower rate than acetoin is oxidised, and it cannot therefore be an intermediate in the reaction. This rules out the possibility of reaction path (a) above; reaction path (b) is evidently the one followed by the V^V oxidation of acetoin:-

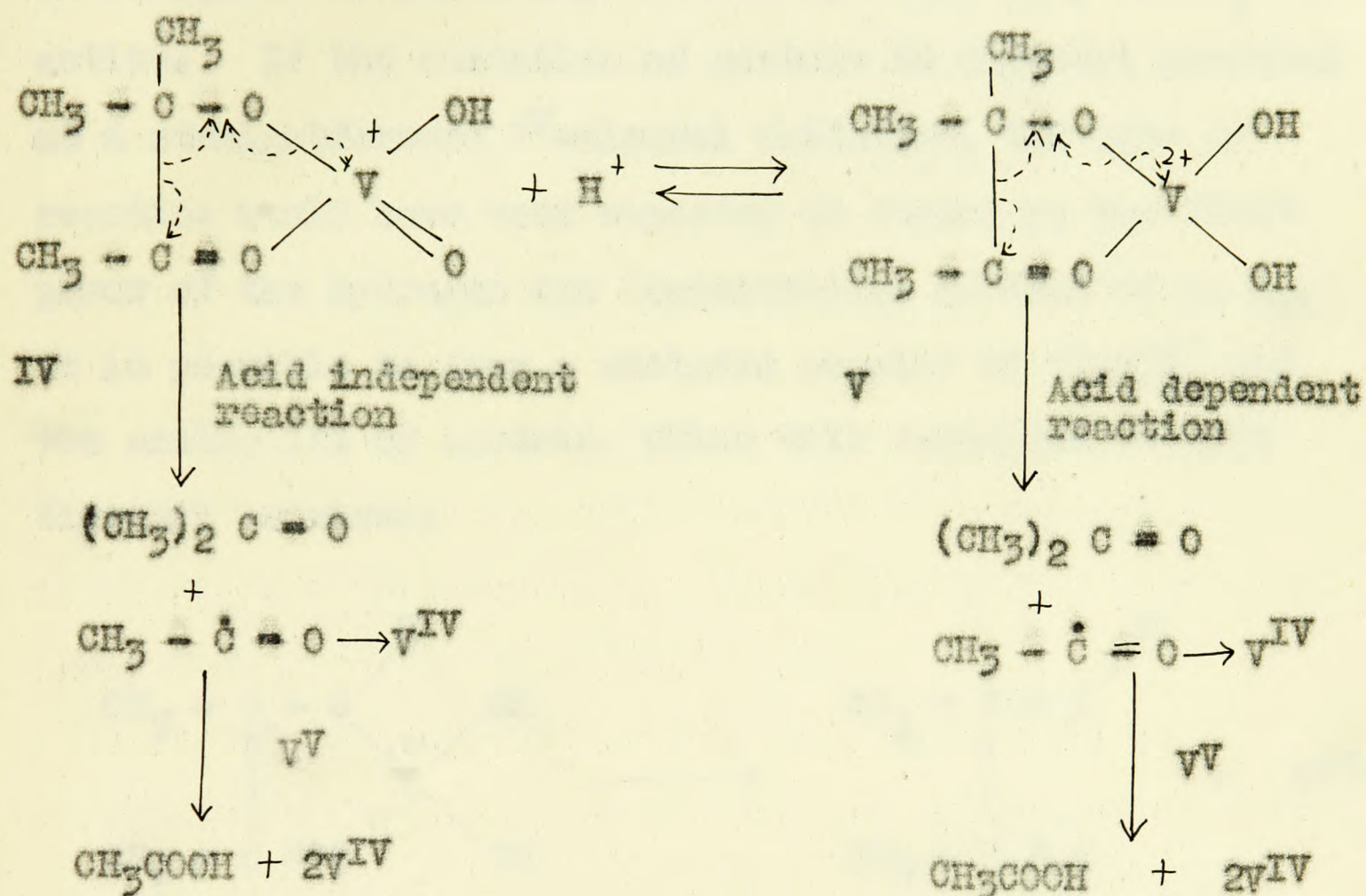


Methyl acetoin only consumes 2.0 equivalents of V^V in its oxidation and must undergo direct C - C bond fission.



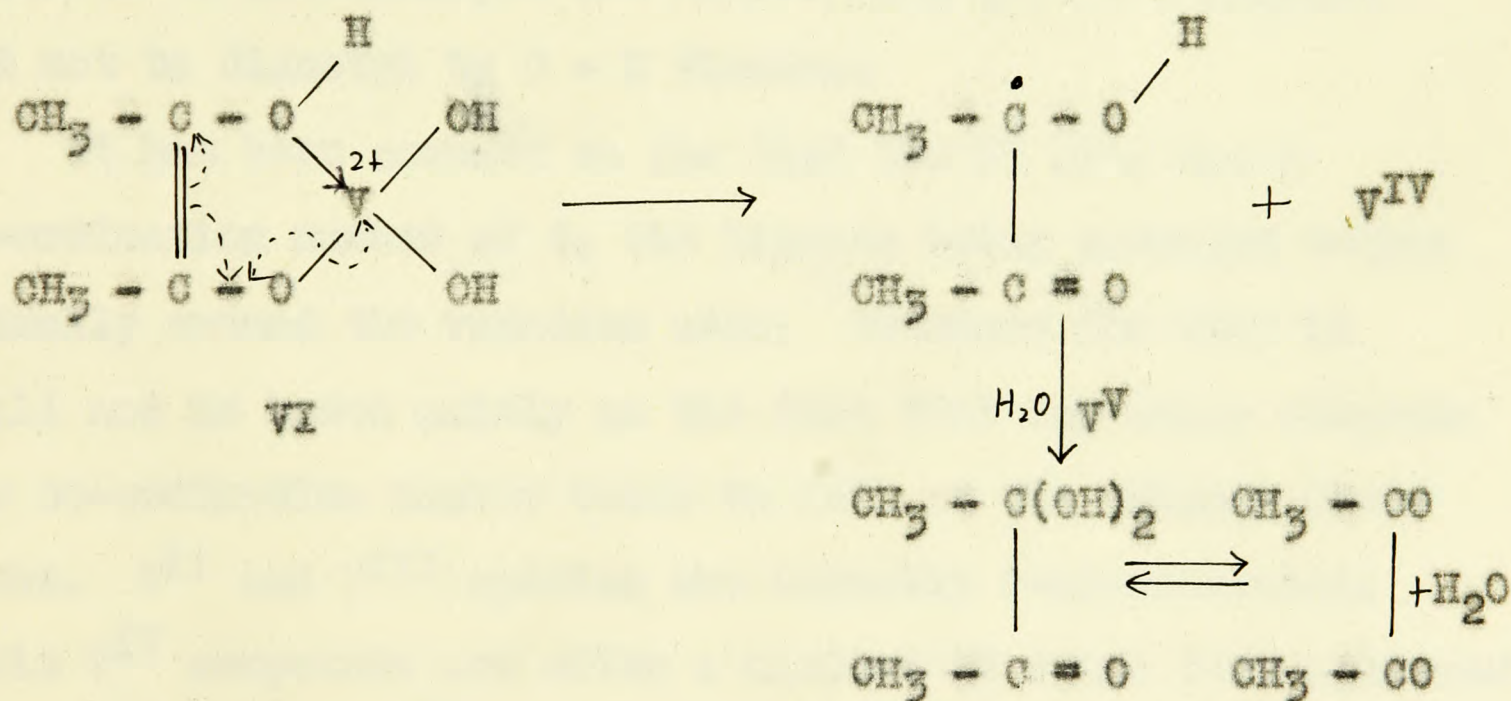
The kinetics of the methyl acetoin oxidation are similar

to those described in Chapter 4 for the hydroxy acids and for pinacol. A reaction mechanism similar to that proposed for the hydroxy acids can be written for this oxidation:-



Exactly the same arguments can be applied in favour of this mechanism that were used in Chapter 4 when the mechanism of the oxidation of hydroxy acids was being discussed. There seems to be no reason to doubt that, if the mechanism proposed for the hydroxy acid oxidation is correct, then the above mechanism operates in the oxidation of this hydroxy ketone, with complexes IV and V being present in small concentrations only.

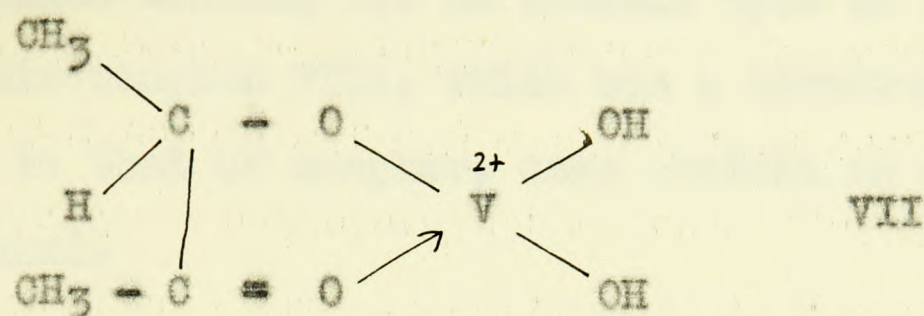
The acid dependence of the acetoin oxidation, however, is different from any yet observed in a V^V oxidation. The absence of an acid independent reaction indicates that VO_2^+ is incapable of oxidising acetoin and that only $V(OH)_3^{2+}$ is active. If the oxidation of acetoin to diacetyl proceeded as a straightforward V^V -alcohol oxidation, the rate of reaction would have been expected to depend on the first power of the hydrogen ion concentration instead of on h_0 . It is possible to draw a chelated complex of $V(OH)_3^{2+}$ and the enol III of acetoin, which will decompose to give diacetyl as shown.



However it has been shown, as mentioned above, that the rate of enolisation of acetoin (to give III) in acid solutions is negligible. There is no reason why the presence of V^V in the solution should alter this state of affairs,

and in order to form a complex such as VI the C - H bond of the :CHOH group must be broken in some way.

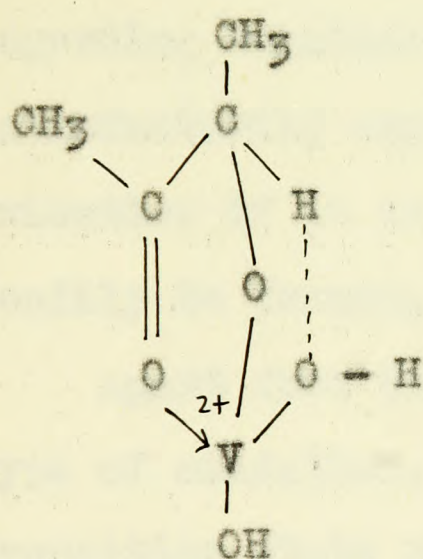
The most likely complex to be formed between $V(OH)_5^{2+}$ and acetoin is



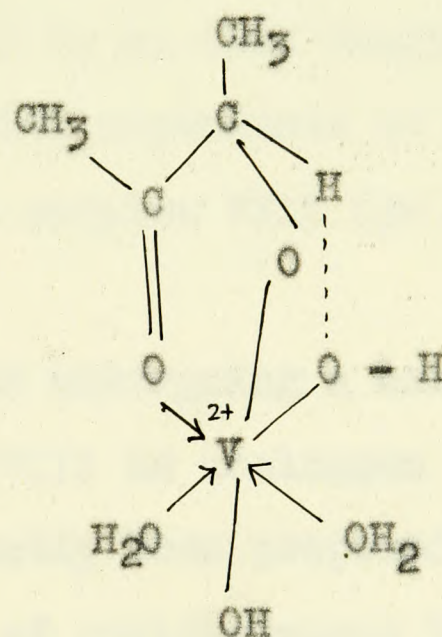
This complex is analogous to that proposed for the acid dependent oxidation of methyl acetoin V, but it would be oxidised to acetaldehyde and acetic acid by C - C fission and not to diacetyl by C - H fission.

It has been assumed so far that the V^V atom has a co-ordination number of 4, the ligands being arranged tetrahedrally around the vanadium atom. Evidence for this is small and is based mainly on the fact that for other elements the co-ordination number tends to fall as the valency state rises. V^{II} and V^{III} species are normally 6-co-ordinated, while V^{IV} compounds are often a mixture of 4 and 6-co-ordinated species⁸⁸. The VO_4^{3-} anion however is known to be tetrahedrally arranged with a V-O bond length of 1.86 Å⁸⁹. In $VOCl_3$ in which compound the vanadium atom is more positively charged, this distance is reduced to 1.56 Å⁰ i.e. rather

longer than C - O or C - C bonds. It is possible to write complex VII in such a way that the H atom of the C - H bond which must be broken in the oxidation, is sufficiently near one of the hydroxyl groups on the V^V atom to give the necessary hydrogen bonding for an alcohol type of oxidation to occur. This complex VIII, which has a structure of similar shape to that of camphor, uses acetoin as a tridentate ligand.



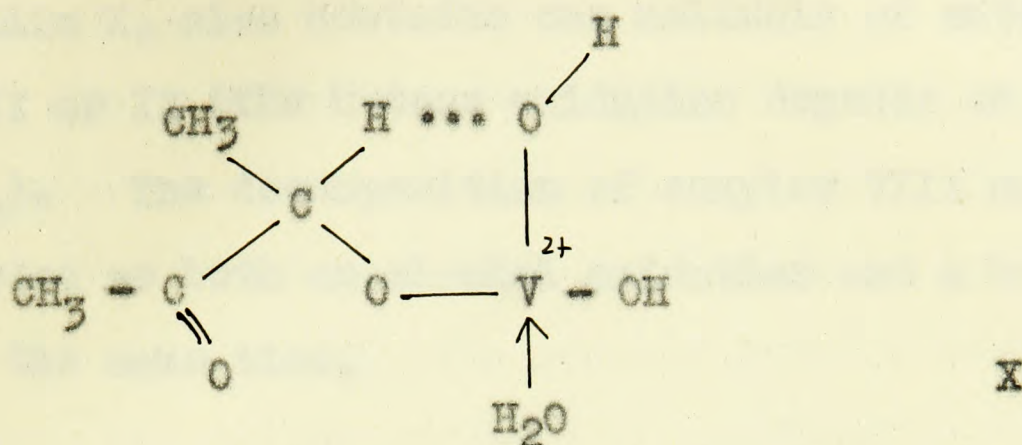
VIII



IX

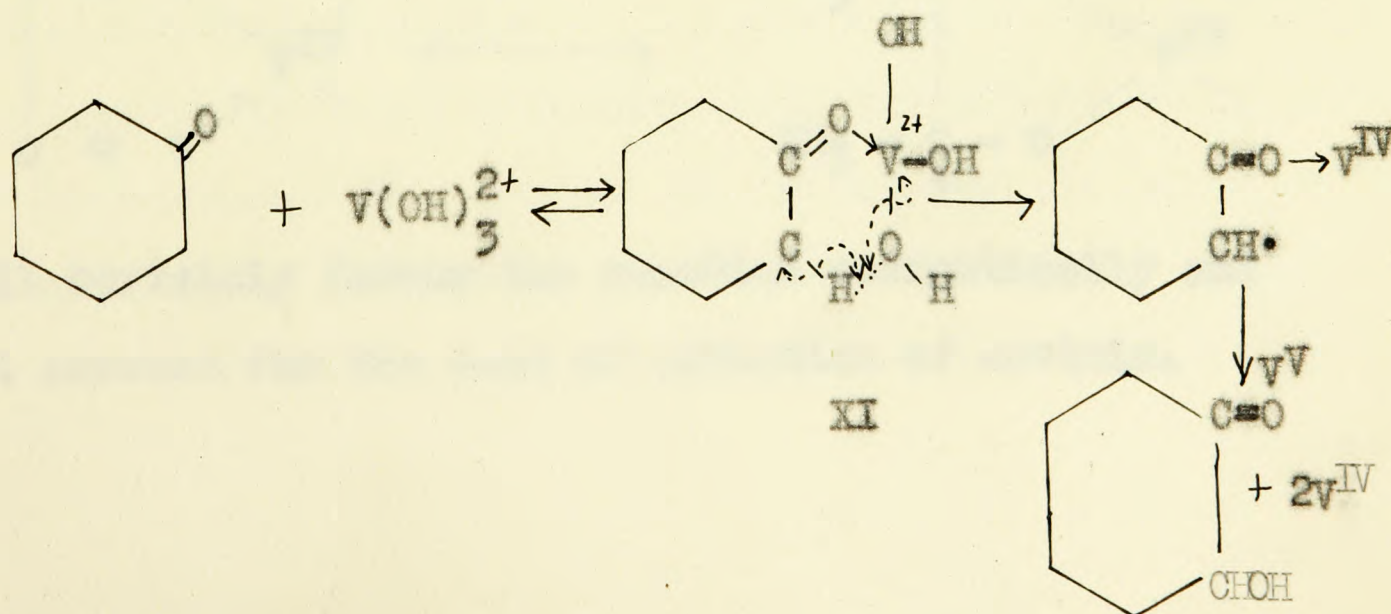
If the V^V atom were octahedrally coordinated, complex IX would result which might be slightly less strained than VIII.

The difference between VIII and complex X which would be expected for an ordinary alcohol type of oxidation

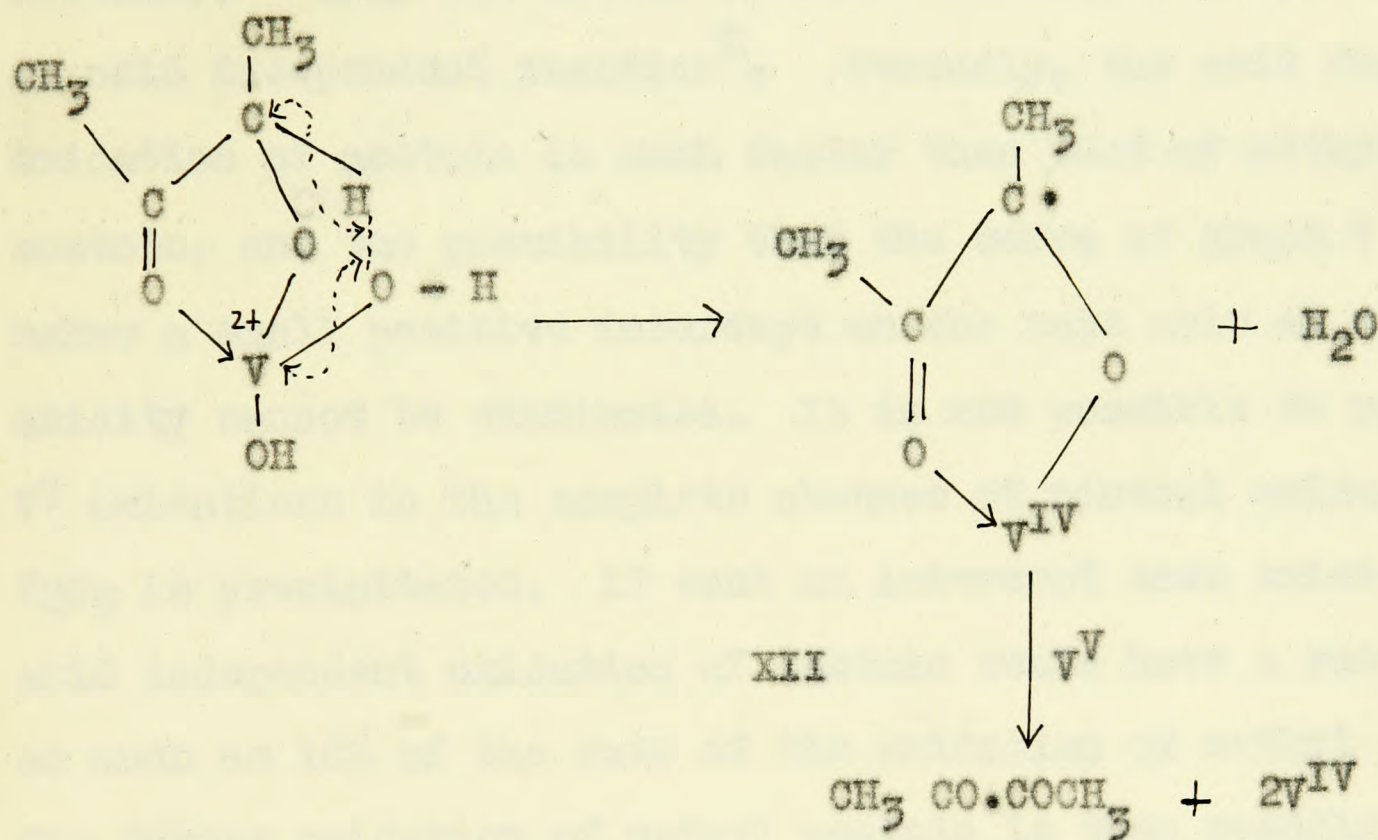


is that X contains an extra molecule of water, which is not present in VIII. The Zucker-Hammett hypothesis regarding reactions which depend on h_0 (see Chapter 2) will satisfactorily explain the acidity dependence of the acetoin oxidation if it is assumed that complex VIII (or IX) can readily be formed.

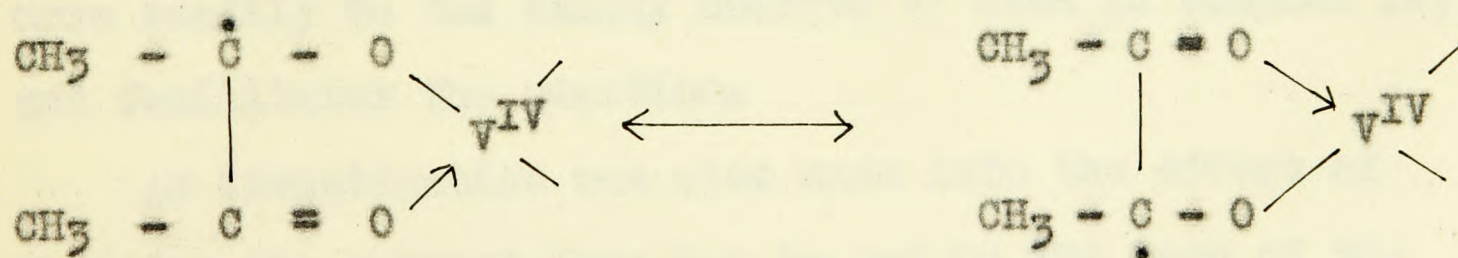
Apart from being capable of undergoing a normal alcohol type of oxidation, the complex VIII is analogous to the transition state which has recently been proposed⁹⁰ for the mechanism of the oxidation of cyclohexanone by V^V , a reaction which proceeds by direct attack of $V(OH)_3^{2+}$ on the ketone rather than on its enolic form:-



Complex XI, like X, also contains one molecule of water more than VIII or IX (The ketone oxidation depends on $[H^+]$ and not on h_o). The decomposition of complex VIII may thus be regarded as both an alcohol oxidation and a ketone oxidation at the same time.



The intermediate V^{IV} complex radical XII can have a mesomeric structure



This will certainly favour the reaction energetically and may well account for the ease of oxidation of acetoin.

It remains to account for the apparent inability of the ion VO_2^+ to attack acetoin. It should first be noticed that the situation is comparable with that of the oxidation of the glycols 2.3 butane diol (the reduced form of acetoin) and 2 methyl 2.3 butane diol (the reduced form of methyl acetoin). Only the latter of these is oxidised by V^{V} in an acid independent reaction⁵. Secondly, the acid dependent oxidation of acetoin is much faster than that of methyl acetoin, and the possibility that the curve of graph 7 (a) makes a small positive intercept on the rate axis at zero acidity cannot be eliminated. It is not possible to perform V^{V} oxidations in the complete absence of mineral acids, as V_2O_5 is precipitated. If such an intercept does exist the acid independent oxidation of acetoin could have a rate of as much as 10% of the rate of the oxidation of methyl acetoin. The faster oxidation of methyl acetoin is then readily explained by the inductive electron releasing effect of the extra methyl group which enables electrons to be transferred more readily to the singly charged V^{V} atom in complex IV, and facilitates the reaction.

An investigation was also made into the effect of changing the solvent from H_2O to D_2O in the case of the acetoin oxidation. In 0.38 M solutions of sulphuric acid, the mean of the results of several experiments showed that

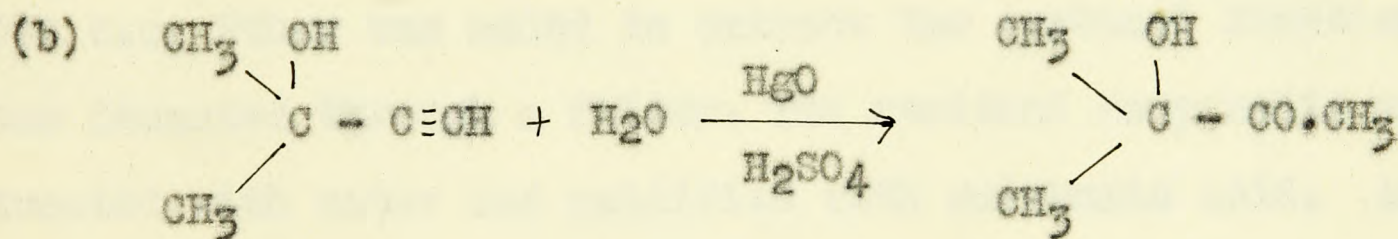
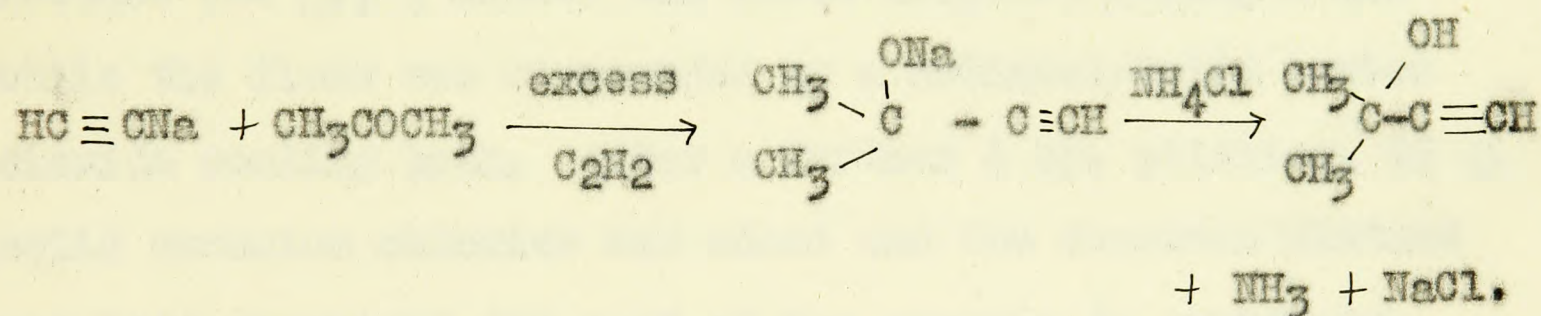
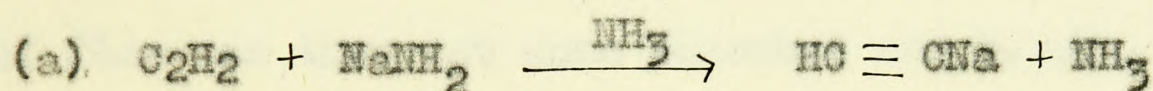
the reaction proceeded faster in heavy water by a factor of 1.3. For an acid catalysed reaction a solvent isotope effect k_{D_2O}/k_{H_2O} of about 2 is to be expected. This leaves an isotope effect k_{H_2O}/k_{D_2O} of about 1.5 to be accounted for by some other means. This ratio is too small for a primary isotope effect and must be accounted for as a secondary isotope effect. The only enol form of acetoin which can form in neutral or acid solutions is $CH_3 \cdot CHOH \cdot C(OH) : CH_2$. The species formed by leaving acetoin in heavy water will therefore be $CH_3 \cdot CHOH \cdot CO \cdot CD_3$ and it must be the effect of the CD_3 group which produces the secondary isotope effect concerned. The effect of replacing a CH_3 group of the radical complex XII by CD_3 would be to reduce the rate of formation of the species slightly. This would result in a decrease in the rate of the overall reaction and would account for the observed secondary isotope effect.

Experimental

Acetoin: A commercial product (B.D.H.) was washed thoroughly with ether to remove diacetyl and was sucked dry at the pump. The resulting white crystals of the dimer melted on rapid heating at 125° . Analysis of the compound gave:- Found: C, 54.67; H, 9.16, Calc. for $C_4 H_8 O_2$.
C, 54.59; H, 9.16%.

A standard solution of the material was made up by dissolving a weighed amount of the solid in distilled water, which had previously been boiled out and saturated with nitrogen. The solution was always kept under an atmosphere of nitrogen in order to prevent aerial oxidation to diacetyl occurring.

Preparation of 3-Hydroxy 3-methyl 2-butanone.



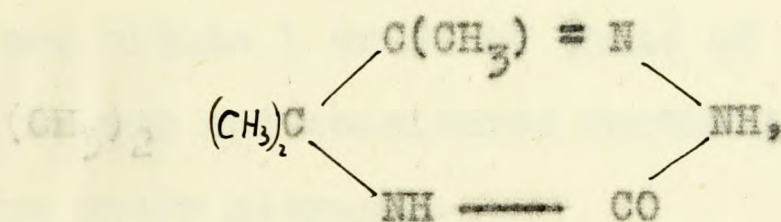
(a) Preparation of Dimethyl Ethynyl Carbinol^{91,92,93.}

Sodium (17½ g., ¾ mole) was added in small pieces during ½ hr. to 750 c.c. liquid ammonia contained in a 1 l. 3-necked round bottom flask fitted with a mechanical stirrer. Once the reaction had started nitrogen was bubbled through the solution. After stirring for a further 1½ hr. acetylene was bubbled through the mixture in place of nitrogen. The

acetylene was passed through an apparatus consisting of a mercury safety valve, an empty Dreschel bottle cooled in a mixture of solid carbon dioxide and methanol, a Dreschel bottle containing sulphuric acid and a third, reversed, Dreschel bottle, before entering the reaction system. The first two bottles were intended to absorb any acetone which should splash out of the acetylene cylinder. The acetylene was passed for $1\frac{1}{2}$ hr. at a rate of 1 l. a minute. The rate of flow was then cut down somewhat and B.D.H. "AnalaR acetone (44 g., $\frac{3}{4}$ mole.) was added dropwise during 1 hr. while the flask was surrounded by a methanol/solid carbon dioxide cooling bath. After a further 4 hr. stirring, 45 g. solid ammonium chloride was added and the reaction mixture was left overnight for most of the ammonia to evaporate. 300 c.c. Ether was added to extract the carbinol formed and was decanted through a filter. The residual damp solid was treated with water and acidified with sulphuric acid. An oily layer which separated was extracted into ether and combined with the first ether extract. Most of the ether was distilled to drive off residual ammonia and the remaining material was dried over anhydrous potassium carbonate after adding some fresh ether. After removal of ether, the product distilled at 101-102°. Yield = 39.0 g (62%), $n_D^{20} = 1.421$ (reported⁹² 1.4211).

(b) Preparation of 3-Hydroxy 3-methyl 2-butanone⁹⁴

Dimethyl ethynyl carbinol (40 g.) was added gradually to 125 c.c. water, 10 c.c. sulphuric acid and 2 g. yellow mercuric oxide in a 250 c.c. round bottomed flask. The mixture was steam distilled and the distillate was saturated with sodium chloride to separate the product. The product was extracted into ether, dried over magnesium sulphate and fractionally distilled. Yield of material of B. Pt. 136-138° was 19.4 g. $n_{20}^D = 1.4150$ (reported⁹⁴ 1.4150). A semicarbazone was prepared and recrystallised from ethanol to M. Pt. 164° (reported⁹⁶ 164°). A material of melting point 210° was also obtained from the semicarbazone which gave an analysis close to that required by the compound



which would be formed by dehydration of the semicarbazone.

A standard solution of the hydroxy ketone was made up in the same way as that used in the case of acetoin.

Procedure for Studying the Reactions

This was basically the same as that described in Chapter 4 for the hydroxy acid oxidations. Some difficulty was experienced at first in observing the end point of the titration. It appeared that there was a tendency for the indicator to oxidise acetoin which was present in excess in solution. It was found that a good end point could still be obtained if one or two extra drops of indicator were added near the end of the titration and if the phosphoric acid concentration was kept close to that mentioned in Chapter 4.

In view of the ready oxidation of acetoin by air, some of the preliminary reactions were performed under nitrogen in the manner previously described. However the observed rates were within 1 or 2% of those of reactions performed in air and it was not considered worthwhile to do the other reactions under nitrogen.

Reactions in Heavy Water. Deuterium oxide was obtained in 99.8% purity from Norsk Hydro. The reaction involving heavy water as solvent and those using light water as a comparison were followed by measuring the increase in the vanadium (IV) absorption at $750\text{ m}\mu$ in a "Unicam" SP 500 spectrophotometer. 2 Mm. quartz cells were used, and the temperature was maintained at $25^{\circ} \pm 0.1^{\circ}$ by means of an Adkins thermostatted

cell holder. Reaction solutions were made up by using an "Aglar" micrometer syringe, the organic substrate being added last. The mixture was shaken thoroughly at room temperature to obtain homogeneity and was then quickly transferred to the quartz cell by means of a second syringe. Two to three minutes were usually required for the mixture to reach the temperature of the thermostat.

The vanadium (V) solution which was used for the oxidations in heavy water was prepared from solid ammonium metavanadate, heavy water and concentrated sulphuric acid (H_2SO_4); the final water was therefore 98.8% D_2O . The vanadium solution was found to lose 5 - 10% of its strength while being measured out as it attacked the steel needle of the syringe.

The acetoin solution was made up by dissolving the solid in heavy water which had been saturated with nitrogen. The solution was kept under nitrogen in a serum capped test tube for 3 days in order to allow any exchange of hydrogen with deuterium to occur.

CHAPTER 7

Oxidation of Some Aldehydes

Some proposed mechanisms of oxidation of aldehydes by various reagents were described in Chapter 2. It will be seen from these that each proposed mechanism is different in some respect from the others, and it seems most unlikely that all of them can be entirely correct. In the case of the chromium (VI) oxidation which has been studied by more than one group of workers there is still controversy as to which mechanism fits the experimental observations best.

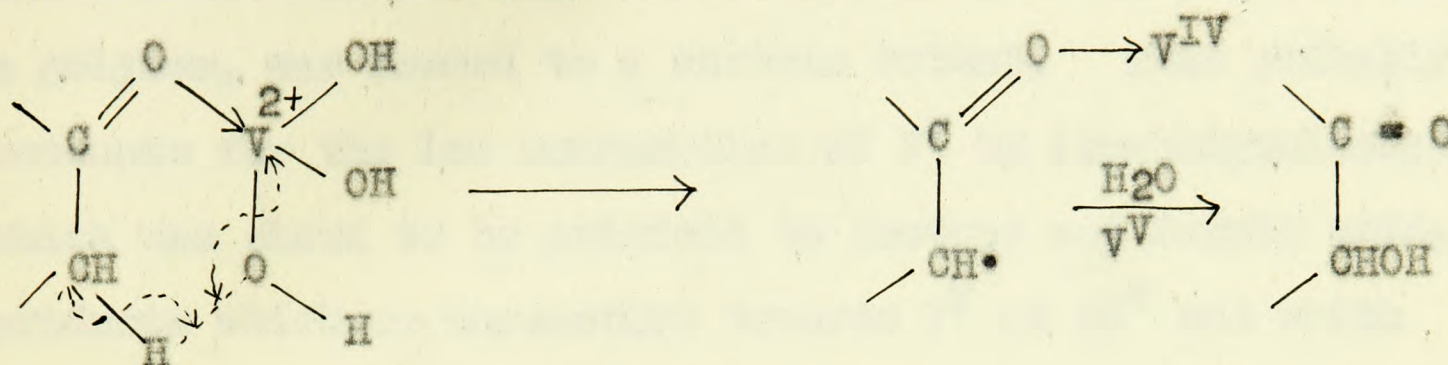
There is some similarity, however, between the mechanisms of oxidations of aldehydes by the one-electron oxidising agents $\text{Fe}(\text{CN})_3^-$, Ce^{IV} , Mn^{III} and by the free radical $\cdot\text{ON}(\text{SO}_3\text{K})_2$, each of which is reported to attack the enolic form of the aldehyde, or the enolate anion. For some simple aldehydes and ketones the rate of oxidation by these reagents approaches a limit and becomes independent of the concentration of oxidant. Where enolisation rates have been measured, the rate of formation of the relevant enol is closely similar to the rate of oxidation. The products of these one-electron oxidations, are those that

would be expected if the oxidant attacked the α position of the aldehyde. The initial product is of the type $R \cdot \text{CHOH} \cdot \text{CHO}$, which undergoes C-C bond fission on further oxidation. Aldehydes such as benzaldehyde, formaldehyde and furfuraldehyde, which cannot enolise normally, are either attacked very slowly, or not at all, by these reagents.

A different mechanism was proposed by Bawn et al⁹⁵ for the cobalt (III) oxidation of acetaldehyde, in which it was suggested that this powerful one-electron reagent extracted a hydrogen atom to give the radical $\text{CH}_3 \text{CO} \cdot$ which leads eventually to acetic acid. No attack at the α carbon atom occurs.

A mechanism has recently been proposed by J.S. Littler⁹⁰ for the oxidation of cyclohexanone by one-electron oxidants. This was briefly mentioned in Chapter 6. It was found that cerium (IV) attacks the ketone at a rate faster than the rate of enolisation, and must be attacking the cyclohexanone in its ketonic form. Solvent isotope effects suggest that vanadium (V) and manganese (III) (uncomplexed and in acid solution) also attack the ketone directly rather than its enol. Littler proposed that a complex was formed between the oxidant and the ketone, and that this decomposed by a cyclic electron movement involving an adjacent C-H bond;

e.g. for the V^V oxidation



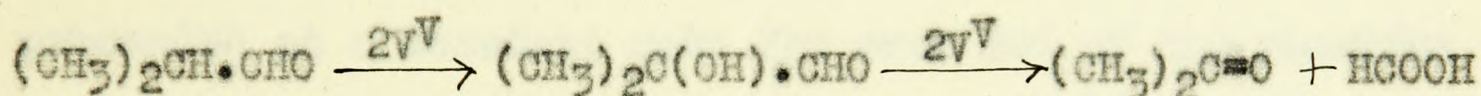
This mechanism could also apply to aldehydes when attack at an α C-H bond is involved.

The aldehydes propionaldehyde, n-butyraldehyde and iso-butyraldehyde are readily oxidised by V^V and are sufficiently soluble in water for their oxidations to be investigated in aqueous solutions at 25° . Acetaldehyde was not used as it is not attacked by V^V at 25° and it volatilises from solution at temperatures greater than this. It was therefore decided to examine the oxidation kinetics in some detail and determine whether or not oxidation occurred through the enol, and if not, whether there was any similarity to the ketone oxidation.

Results

The three aldehydes which were examined consumed V^V readily at 25° . In 1 M sulphuric acid the maximum amounts of V^V reduced by 1 mole of aldehyde were propionaldehyde

3.9 moles, n-butyraldehyde 8.1 moles, iso-butyraldehyde 2.7 moles. In all cases an insoluble material, presumably a polymer, was formed to a certain extent. This probably accounts for the low consumption of V^V by iso-butyraldehyde which was shown to be oxidised to acetone and formic acid, products which are unreactive towards V^V at 25° and which require 4.0 equivalents for complete formation. Thus the reaction is



A small amount of an unidentified acid product of empirical formula $\text{C}_8\text{H}_{14}\text{O}_3$ was also obtained from the iso-butyraldehyde oxidation. The absorption of 4 equivalents of V^V by propionaldehyde is sufficient for oxidation to acetaldehyde and formic acid, while n-butyraldehyde requires 8 equivalents of V^V for oxidation to 1 mole of acetaldehyde and 2 moles of formic acid. Other carboxylic acids such as propionic acid or butyric acid cannot be formed intermediately as these are not attacked by V^V under the conditions used. Clearly the oxidations involve attack at the α C-H group of all three aldehydes.

The precipitation of insoluble polymer occurred to some extent in all the oxidations which were followed titrimetrically, but fortunately the polymer did not appear

immediately on mixing the reactants, and for propionaldehyde and n-butyraldehyde good first order plots were obtained for 50% of the disappearance of V^V after a short initial induction period. After this the rate of reaction tended to fall off from a first order reaction to a certain extent depending on the conditions used. In the case of iso-butyraldehyde, however, the rate of disappearance of V^V was independent of the concentration of V^V and the reaction proceeded at a constant rate for over 90% of the reaction if the polymer was not formed to an appreciable extent.

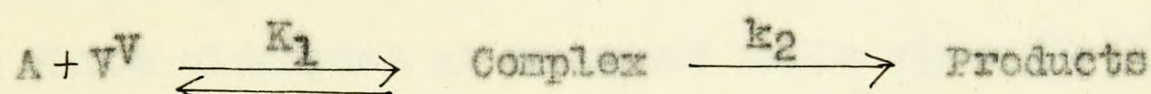
Some reactions were carried out under nitrogen, but these proceeded at the same rate as reactions performed under similar conditions in air. Polymer was still formed in the reactions under nitrogen, but did not appear in the solutions which did not contain V^V , which were used for presaturating the nitrogen. It appeared that the V^V oxidations initiated polymerisation, possibly through the formation of free radical intermediates.

Table II shows that the initial rate of oxidation of iso-butyraldehyde is nearly proportional to the concentration of aldehyde. The 10% drop in the value of the rate constant is very small for a 5-fold increase in the concentration of aldehyde. A much more marked tendency for the first order rate constant to fall was noted for the other two

aldehydes and is clearly shown in the fifth columns of Tables 12(a)-(d).

In the oxidation of ethanol by cerium (IV) Ardon⁹⁶ noticed a similar effect when the alcohol concentration was changed. Littler⁹⁷ also found this effect in the cerium (IV) oxidation of cyclohexanol. Ardon explained his results on the basis of the extensive formation of a complex between the reactants.

In the present case, if A = aldehyde $[V^V]$ = concentration of uncomplexed V^V , $[V_T^V]$ = concentration of total V^V .



$$\text{Rate of oxidation} = \frac{-d[V_T^V]}{dt} = k_2 [\text{Complex}]$$

$$\text{Now } [\text{Complex}] = K_1 [V^V][A] \quad (K_1 = \text{Equilibrium constant}).$$

$$\text{and } [V^V] = [V_T^V] - [\text{Complex}]$$

$$\text{Therefore } [\text{Complex}] \cdot (1 + K_1 [A]) = K_1 [V_T^V][A]$$

$$\text{Therefore } \frac{-d[V_T^V]}{dt} = \frac{k_2 K_1 [V_T^V][A]}{1 + K_1 [A]}$$

$$\text{or } \frac{-1}{\frac{d[V^V]}{dt}} = \frac{a}{[A]} + b \quad \text{where } a \text{ and } b \text{ are constants.}$$

When the concentration of complex is very small, b will also be small and vice versa.

A graph of $\frac{1}{\text{Rate}}$ against $\frac{1}{[A]}$ should then give a straight line of slope a . The final columns of Tables 12 (a) (c) and (d) show that this relationship holds well for propionaldehyde at 1.0 M acid and for n-butyraldehyde at 1.0 and 3.0 M acid, but table 12 (b) shows that for propionaldehyde in 3.0 M acid the relationship only holds at low aldehyde concentrations.

The dependences of the rates of oxidation of the three aldehydes on acidity are also somewhat complex. Graph 10 shows the general trend of the initial rates of reaction of the three aldehydes at constant ionic strength (5.1 M). All three curves tend to straight lines passing through the origin at low acidities (0.5 - 2.5 M), but positive deviations from this line occur at higher acidities. The results do not fit the Hammett acidity function h_0 since the portions of the curves at low acidities follow hydrogen ion concentration too closely. The rate of oxidation of propionaldehyde is nearly proportional to $[H^+]^2$ as can be seen from Graph 11.

Graph 12 shows the acidity dependence of the initial rate of reaction of iso-butyraldehyde with 0.05 N solutions of vanadium (V), iodine and bromine. The rates of the iodine and bromine reactions are not strictly independent

of halogen concentration because the reaction of iodine with an enol is a reversible process and both reactions involve some oxidation of the aldehyde as well as addition to the double bond. The oxidation reaction is of first order with respect to halogen concentration, while the addition reaction is of zero order. The three curves are accurately parallel above 1.0 M acid concentration, and if it is assumed that the rate of oxidation by the halogens is independent of acid concentration in this range, then it follows that the vanadium (V) oxidation occurs at the same rate as the iodine and bromine addition to the enol. This is conclusive evidence that the vanadium (V) is attacking the enol form of the aldehyde, and that the acid catalysed enolisation of the aldehyde is the rate determining step in the oxidation. The slight deviation of the rate of reaction/acidity curve from linearity at higher acidities is not inconsistent with enolisation of the aldehyde. The acid dependence of the rate of enolisation of acetophenone which was measured by Zucker and Hammett³⁷ only showed an accurate proportionality to acid concentration at low acidities.

An attempt was made to measure the rates of enolisation of the other two aldehydes by their reactions with bromine and iodine. In the case of the iodine reaction the

equilibrium was too much in favour of the aldehyde to facilitate accurate measurements of the rate of reaction. Bromine again gave a considerable amount of oxidation reaction, but it was possible to perform a rough analysis of the graph of concentration of bromine against time into its first and zero order components in bromine concentration. Taking the zero order component as the rate of enolisation it was found that this was 4 to 5 times as large as the rate of the vanadium oxidation for both aldehydes for 0.05 M solution of V^{V} . This value was not appreciably altered by changing the acidity of the solution from 1.0 M to 3.0 M.

Table 11. Oxidation of iso-butyraldehyde with vanadium (V)

Dependence of zero order rate constants on aldehyde concentration at 25°C; 1.0 M perchloric acid, Ionic strength = 3.1 M

[Aldehyde] (M)	(Moles $\frac{10^6 k_0}{V^V}$, sec ⁻¹ , litre ⁻¹)	$\frac{10^6 k_0}{[Aldehyde]}$
0.088	0.84	9.5
0.176	1.63	9.3
0.220	1.95	8.9
0.440	3.75	8.5

Table 12. Oxidation of propionaldehyde and n-butyraldehyde with vanadium (V) Dependence of rate of reaction on aldehyde concentration.

$$\frac{1}{k_1} = \frac{a}{[\text{aldehyde}]} + b \quad \text{First order rate constants } k_1$$

(a) Propionaldehyde; 1.0 M perchloric acid; Ionic strength = 3.1 M; Temp. = 25°C.

[Aldehyde] (M)	$\frac{1}{[\text{Aldehyde}]}$	$10^5 k_1 (\text{sec}^{-1})$	$\frac{1}{10^3 k_1}$	$\frac{10^5 k_1}{[\text{aldehyde}]}$	$\frac{10^{-3} a}{(2.60 \times 10^3)}$ (when $b = 2.60 \times 10^3$)
0.069	14.49	3.37	29.7	48.8	1.87
0.104	9.62	4.68	21.4	45.1	1.95
0.138	7.25	6.53	15.3	47.3	1.75
0.176	5.68	7.88	12.69	44.8	1.77
0.220	4.55	9.25	10.81	42.0	1.80
0.264	3.79	10.60	9.43	40.2	1.80
0.308	3.25	11.50	8.70	37.3	1.87
0.352	2.84	13.00	7.69	36.9	1.79
0.460	2.27	15.08	6.70	34.3	1.80

(b) Propionaldehyde; 3.0M perchloric acid, Ionic Strength =
3.1 M, Temp. = 25°C.

$\frac{[\text{Aldehyde}]}{[\text{Aldehyde}]}$	$\frac{1}{[\text{Aldehyde}]}$	$10^5 k_1 (\text{sec}^{-1})$	$\frac{1}{10^3 k_1}$	$\frac{10^5 k_1}{[\text{aldehyde}]}$	$\frac{10^{-3} a}{(b_3 = 1.50 \times 10^3)}$
0.069	14.49	9.85	10.15	142	0.597
0.104	9.62	13.7	7.30	132	0.593
0.138	7.25	17.0	5.88	123	0.604
0.176	5.68	20.2	4.95	115	0.607
0.264	3.79	28.7	3.48	109	0.522
0.352	2.84	36.7	2.73	104	0.433

(c) n-Butyraldehyde; 1.0M perchloric acid, Ionic strength =
3.1 M, Temp. = 25°C

$\frac{[\text{Aldehyde}]}{[\text{Aldehyde}]}$	$\frac{1}{[\text{Aldehyde}]}$	$10^5 k_1 (\text{sec}^{-1})$	$\frac{1}{10^3 k_1}$	$\frac{10^5 k_1}{[\text{aldehyde}]}$	$\frac{10^{-3} a}{(b_3 = 3.00 \times 10^3)}$
0.088	11.4	3.50	28.6	39.8	2.24
0.131	7.63	4.92	20.3	37.5	2.27
0.175	5.71	6.33	15.8	36.2	2.24
0.219	4.57	7.42	13.5	33.9	2.29

(d) n-Butyraldehyde; 3.0 M perchloric acid, Ionic

Strength = 3.1 M, Temp. = 25°C

$[\text{Aldehyde}]$ (M)	$\frac{1}{[\text{Aldehyde}]}$	$10^5 k_1$	$\frac{1}{10^5 k_1}$	$\frac{10^5 k_1}{[\text{Aldehyde}]}$	$10^{-3} a$ (when $b = 1.00 \times 10^3$)
0.083	11.4	9.33	10.7	106	0.850
0.131	7.63	13.2	7.58	101	0.863
0.175	5.71	17.1	5.85	97.7	0.849
0.219	4.57	20.3	4.92	92.7	0.858

Discussion of Results

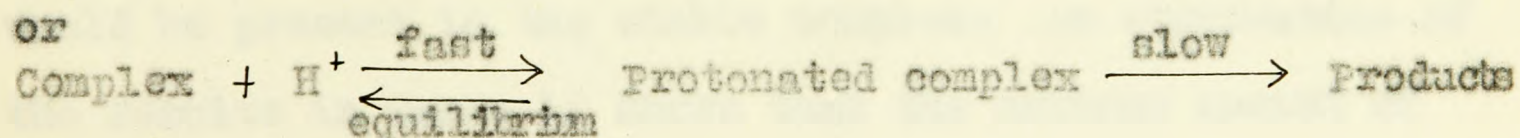
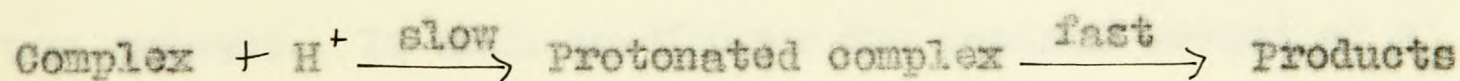
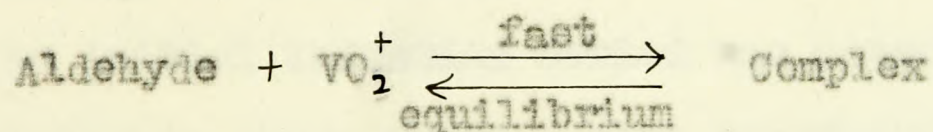
The rate equation for the oxidation of iso-butyraldehyde is $-\frac{d[V^V]}{dt} = k [\text{Aldehyde}] [H^+]$. The rate is independent of V^V . These kinetics are those that would be expected for the enolisation of the aldehyde. As far as can be determined, the rate of oxidation is the same as the rate of enolisation and it is clear that the oxidation of iso-butyraldehyde is occurring by attack of VO_2^+ on the enol at a rate faster than the rate of enolisation. It seems unlikely that $V(OH)_2^{2+}$ is the main oxidising agent, since this would only be present to a small extent in solutions of low acidity and could not be expected to react with all the enol as soon as it is formed. The absence of any isobutyric acid in the products is consistent with the view that V^V attacks the enol form of the

aldehyde.

At low acidities the observed rate equation for the oxidation of both propionaldehyde and n-butyraldehyde is

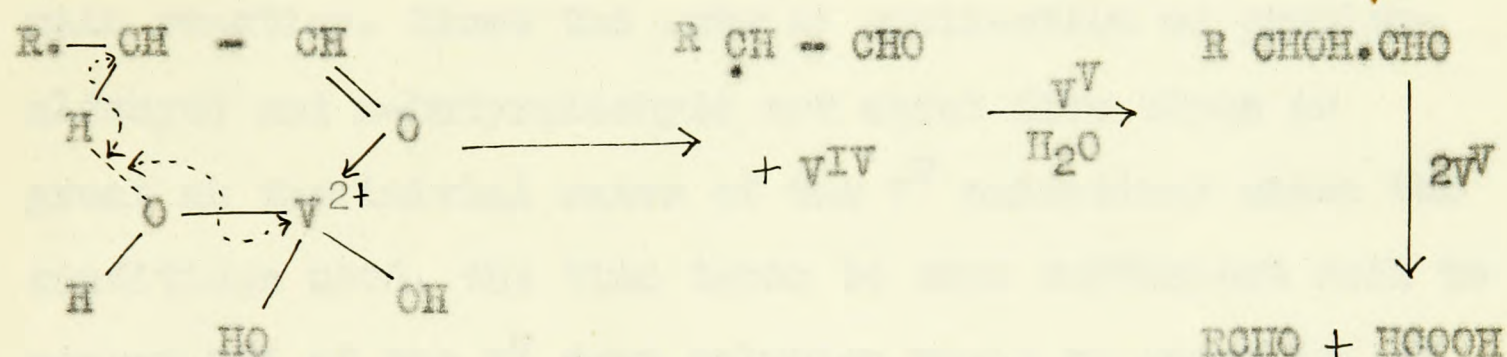
$$\frac{-d[V^V]}{dt} = \frac{k [\text{Aldehyde}][H^+][V^V]}{1 + [\text{Aldehyde}]}$$

The term $(1 + [\text{Aldehyde}])^{-1}$ has already been explained in terms of complex formation between V^V and the aldehyde. The complex which eventually decomposes to products must contain a hydrogen ion as well as a molecule of aldehyde and a VO_2^+ ion, but this protonation may occur either directly on the VO_2^+ - aldehyde complex or previously on the VO_2^+ ion, giving $V(OH)_3^{2+}$ as the complexing species. However it seems unlikely that $V(OH)_3^{2+}$ is formed in sufficient quantity to give enough complex to account for the observed kinetics. The most likely reaction schemes are



The observed kinetics do not show whether the aldehyde is reacting in its normal form or as an enol. Since the

enol is in equilibrium with the aldehyde it may replace aldehyde in the above scheme without altering the expected reaction kinetics. Although an oxidation through the enol would account for the observed C - C bond fission, direct attack on the aldehyde may also give the observed products if Littler's mechanism for the oxidation of ketones by V^V can operate, e.g.:-



However the kinetics of the aldehyde oxidation are not the same as those of the ketone oxidation and it would be wise to say that the above mechanism is no more than a possibility in this case.

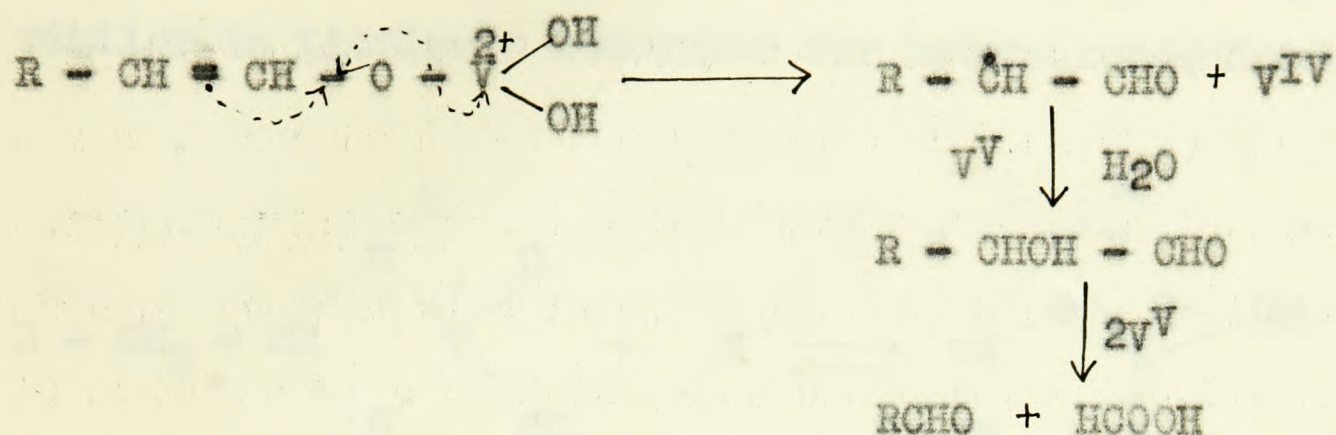
An initial induction period was noted in many reactions, especially in those where a high concentration of aldehyde was present. In such conditions the maximum amount of V^V would be present in the stable complex. An examination of the results in Table 12 shows that the maximum amount of V^V removed as a complex under the conditions used was about 25% with propionaldehyde and 14% with n-butyraldehyde.

If the complex is formed immediately on mixing the aldehyde with V^V there should be no initial induction period to the reaction; but if the complex is formed between VO_2^+ and the enol form of the aldehyde then the initial rate of reaction would depend on the rate at which an equilibrium amount of complex can be formed in solution and an induction period would be observed for the main reaction. Since the rates of enolisation of propionaldehyde and n-butyraldehyde are about five times as great as the initial rates of the V^V oxidations under the conditions used, the time taken to form sufficient enol to remove 25% of the V^V from solution would be about 15% of the half life of the oxidation reaction. If the enol is involved the induction periods of the oxidations should then be of this order. The observed induction periods accord with this predicted value and thus give reason to believe that the complex is formed between VO_2^+ and the enolic form of the aldehydes. However the possibility of slow complex formation between V^V and the aldehyde in its normal form must not be ruled out.

The complex most likely to be formed between VO_2^+ and the enol form of the aldehydes is $R - CH = CH - O - V \begin{smallmatrix} + \\ \nearrow O \\ \searrow OH \end{smallmatrix}$

In order to account for the observed kinetics it must be postulated that this species is stable, but will undergo

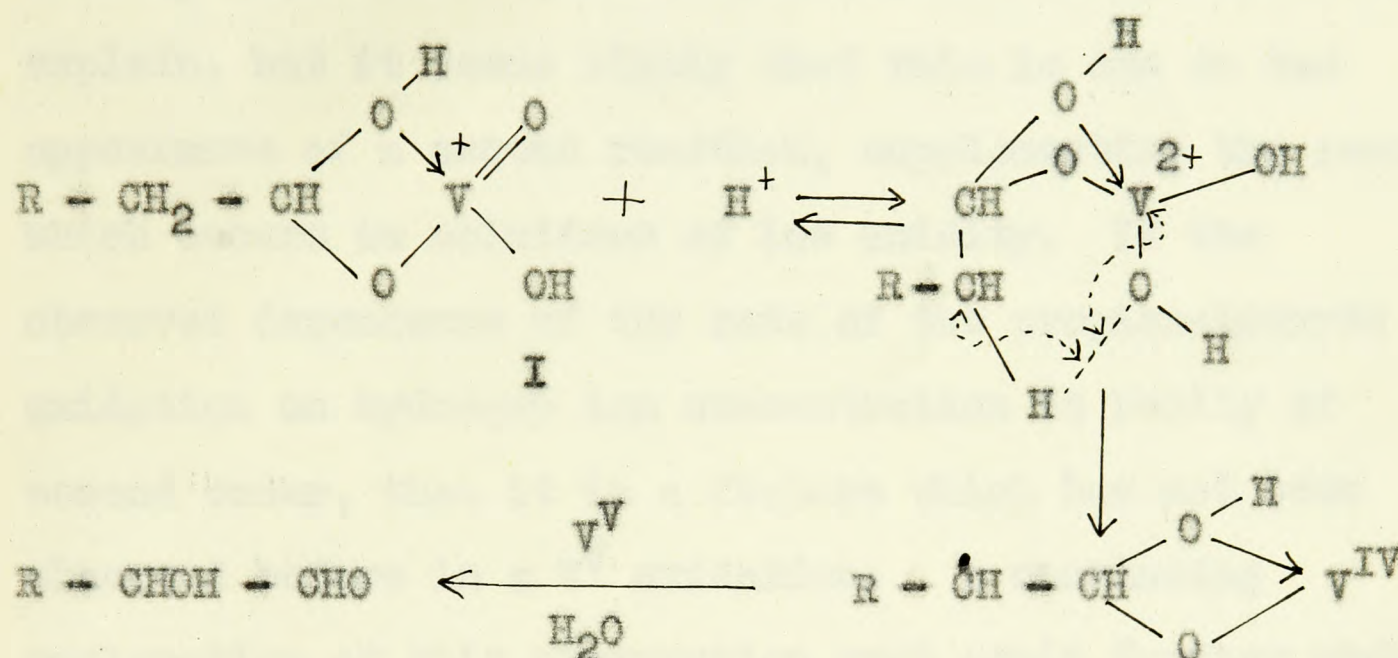
reaction upon protonation:-



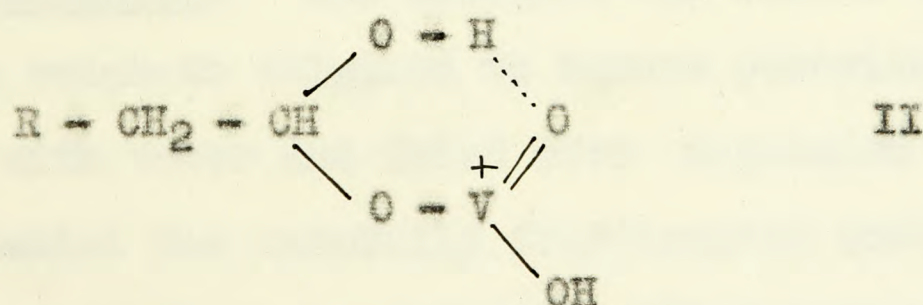
This mechanism must be regarded only as a possibility; in any case it does not account for the observed dependence of rate on acidity at high acid concentrations.

One difference between the structures of aldehydes and ketones such as cyclohexanone is in their respective degrees of hydration. This may have some effect upon the complexing power of the carbonyl group and on the rate of oxidation of these compounds. Cyclohexanone is largely unhydrated in aqueous solutions, but the following figures are given by Hereld⁹⁸ for the degree of hydration of aldehydes: acetaldehyde 66%, propionaldehyde 59%, n-butyraldehyde 55%, iso-butyraldehyde 48%. Bell and Clunie⁹⁹ have more recently measured the degree of hydration of acetaldehyde and obtained a figure of 55%, which is a little less than Hereld's value. It seems likely then that Hereld's figures for the other aldehydes are correct to within about 15%. Thus it is possible that a complex

such as I might be formed between VO_2^+ and the aldehyde, and, after protonation, might decompose by a mechanism similar to Littler's mechanism for ketone oxidations: -



The 4-membered ring involved in this mechanism has no known parallel in other V^{V} complexes, but all possible schemes must be considered in attempting to elucidate a mechanism for any reaction. It is also possible that the complex may take the form II rather than I



but this is more similar to a V^{V} -alcohol complex and after protonation would be expected to decompose to the corresponding acid $\text{R} \cdot \text{CH}_2 \cdot \text{COOH}$ rather than the aldehyde

R CHO.

The much greater dependence of reaction rate on acidity above 3.0 M acid concentrations is difficult to explain, but it seems likely that this is due to the appearance of a second reaction, supplementing the reaction which occurs in solutions of low acidity. If the observed dependence of the rate of the propionaldehyde oxidation on hydrogen ion concentration is really of second order, then it is a feature which has not been observed before in a V^V oxidation. A convincing explanation of this observation must await further work on aldehyde and ketone oxidations both by V^V and by other oxidising agents.

Experimental.

Materials

Propionaldehyde. The aldehyde was shaken with a little ferrous sulphate solution to remove peroxides; it was then washed with water and dried over magnesium sulphate. The material was carefully fractionated under nitrogen and the main fraction was redistilled under nitrogen at 48° .

n-Butyraldehyde and iso-butyraldehyde. These were purified in the same way as propionaldehyde. B. Pts., n-butyraldehyde 73° , iso-butyraldehyde 62° .

Standard Solutions of Aldehydes. These were prepared by dissolving weighed amounts of each aldehyde in distilled water which had previously been boiled out and saturated with nitrogen. The solutions were then made up to a known volume. The standard flasks in which these solutions were prepared were flushed with nitrogen both before use and after each occasion on which the flask was opened later. The strengths of the solutions were checked occasionally by measuring the rate of a standard V^V oxidation of the aldehyde. The rate of reaction was unchanged if the aldehyde solution retained its original strength. Solutions of iso-butyraldehyde showed a slight tendency to lose their original strength; fresh solutions of iso-butyraldehyde were therefore made up from time to time.

Method of Following Reactions.

This was the same as that described in Chapter 4. Some trouble was experienced in detecting the end point of the titration. This was due to reduction of the indicator by the excess aldehyde which was present in the titration mixtures. The trouble was overcome by adopting the precautions described in Chapter 6.

Reactions under nitrogen, and reactions designed to measure the amount of V^V required to oxidise each aldehyde completely, were carried out as described in Chapter 4.

Products of Oxidation.

Each aldehyde was oxidised with a suspension of vanadium (V) oxide in 3 M - 5 M sulphuric acid as described in Chapter 4. The reactions were performed at 25° to minimise the volatilisation of starting materials; reaction was complete in a few hours. The reaction mixture was then fractionated until about a quarter of the total volume had distilled.

The dimedone derivative of acetaldehyde (M. Pt. and mixed M.Pt. 139°) was obtained from the lowest boiling fraction of the propionaldehyde and n-butyraldehyde oxidation mixtures. In the iso-butyraldehyde oxidation a great deal of the polymeric material was formed unless the aldehyde was added gradually to the V^V . Acetone was identified as its 2,4 dinitrophenylhydrazone (M.Pt. and mixed M.Pt. 126°) from the iso-butyraldehyde oxidation. Formic acid was detected by the sensitive chromotropic acid test in fractions boiling at 100° from all three aldehyde oxidations.

Continuous ether extraction of the residue of the oxidation mixtures yielded a little acid material which

was found to be largely formic acid. However small amounts of a white crystalline solid were obtained from the iso-butyraldehyde oxidation. This material gave reactions characteristic of a carboxylic acid. After recrystallisation from water it had M.Pt. 221° .

Found: C, 60.75; H, 8.90; $C_8H_{14}O_3$ requires C, 60.75; H, 8.92%.

Measurement of the Rates of Enolisation of the Aldehydes.

Standard solutions of bromine in sodium bromide and of iodine in sodium iodide were used for measuring rates of enolisation. Halogens were of B.D.H. "AnalaR" grade. The reactions were followed in corked conical flasks at 25° in a similar manner to that in which the V^V oxidations were followed. The reaction mixtures had an initial volume of 100 c.c. and 10 c.c. aliquots were withdrawn at suitable intervals and run into sodium bicarbonate solution to neutralise the acid, and stop the reaction. For the bromine reactions, sodium iodide was also present in the quenching solution to liberate an amount of iodine equivalent to the amount of bromine present. The free iodine was estimated by titration against standard sodium thiosulphate solution.

The reaction of bromine with the three aldehydes is not strictly of zero order with respect to bromine concentration, but also includes a first order oxidation reaction. A rough estimate of the rate of the zero order reaction was obtained as follows and was taken to be the rate of enolisation.

A graph was plotted for each run of the concentration of bromine in the reaction mixture against time. Tangents were drawn to the resulting curve at the beginning of the reaction and at 80% of the reaction as shown in Graph 13. The difference in slopes of these lines then represents $\frac{4}{5}$ of the initial rate of the first order reaction. The rate of the zero order reaction can be calculated by subtracting one quarter of this difference from the slope of the tangent at 80% of the reaction. The value so obtained was checked by measuring the slopes of tangents at other points on the curve (say at 60% and 90% of the reaction).

General Experimental Notes

Melting Points were measured in capillary tubes mounted adjacent to a thermometer in an electrically heated apparatus fitted with a rheostat to vary the rate of heating. The apparatus was calibrated by taking the melting points of various standard specimens from time to time.

Infra Red Spectra were recorded on a Perkin-Elmer Model 136 "Infracord" spectrophotometer using rock-salt plates.

Elemental Analyses were carried out by the Micro-analytical department of the Dyson Perrins Laboratory under the supervision of Drs. Weiller and Strauss.

Thermostat. A lagged thermostat was used for reactions at temperatures above 25°. The temperature was controlled to $\pm 0.1^\circ$ by means of a spiral toluene-mercury regulator connected through a relay box to the heating elements of the thermostat. The water in the thermostat was covered with a layer of paraffin oil to minimise evaporation. A "Towers" thermostat was used for work at 25°. The thermostat was not lagged and the oil layer was unnecessary at this temperature.

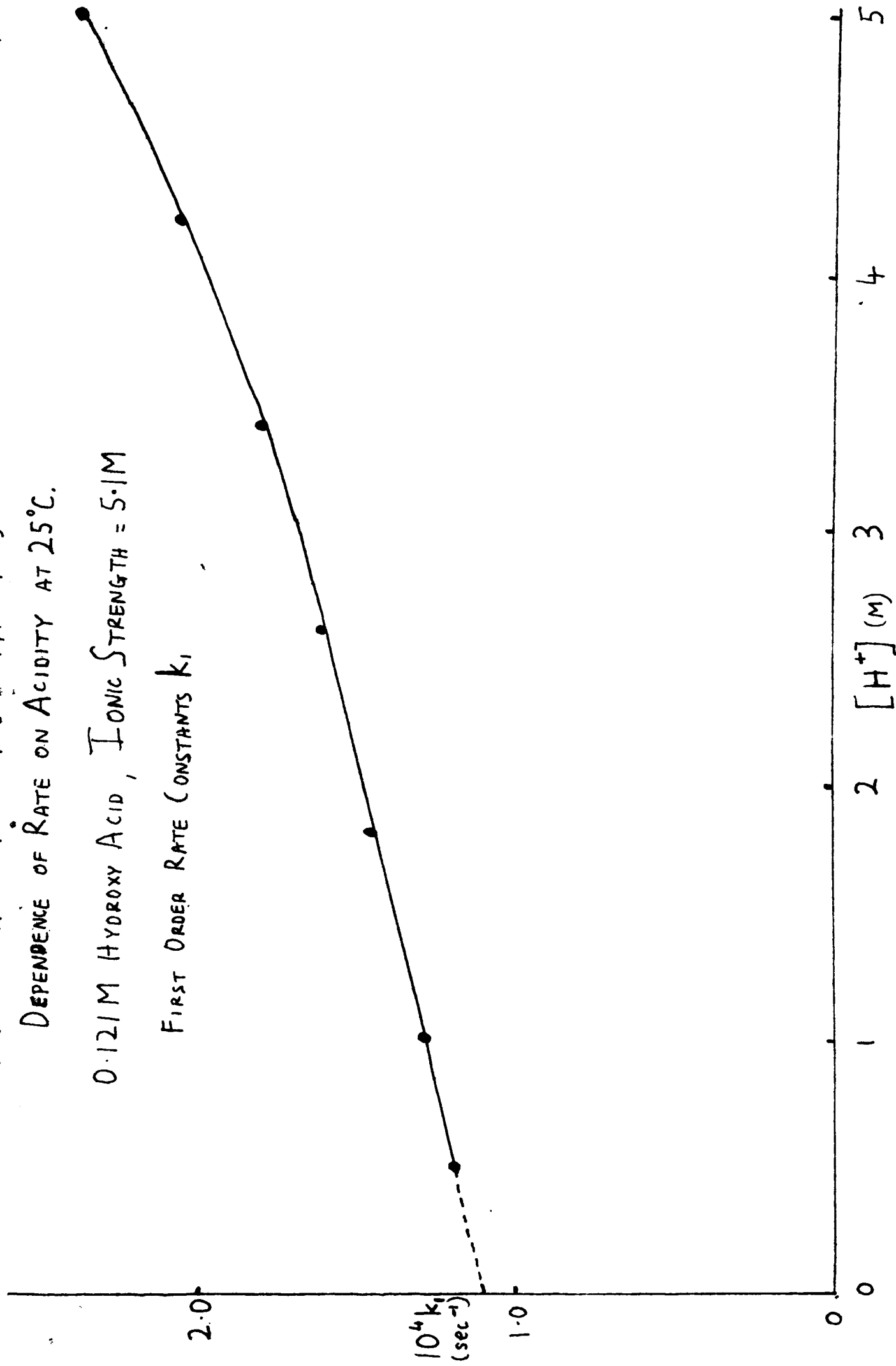
GRAPH 1.

OXIDATION OF α -HYDROXY i-BUTYRIC ACID

DEPENDENCE OF RATE ON ACIDITY AT 25°C.

0.121M HYDROXY ACID, IONIC STRENGTH = 5.1M

FIRST ORDER RATE CONSTANTS k_1



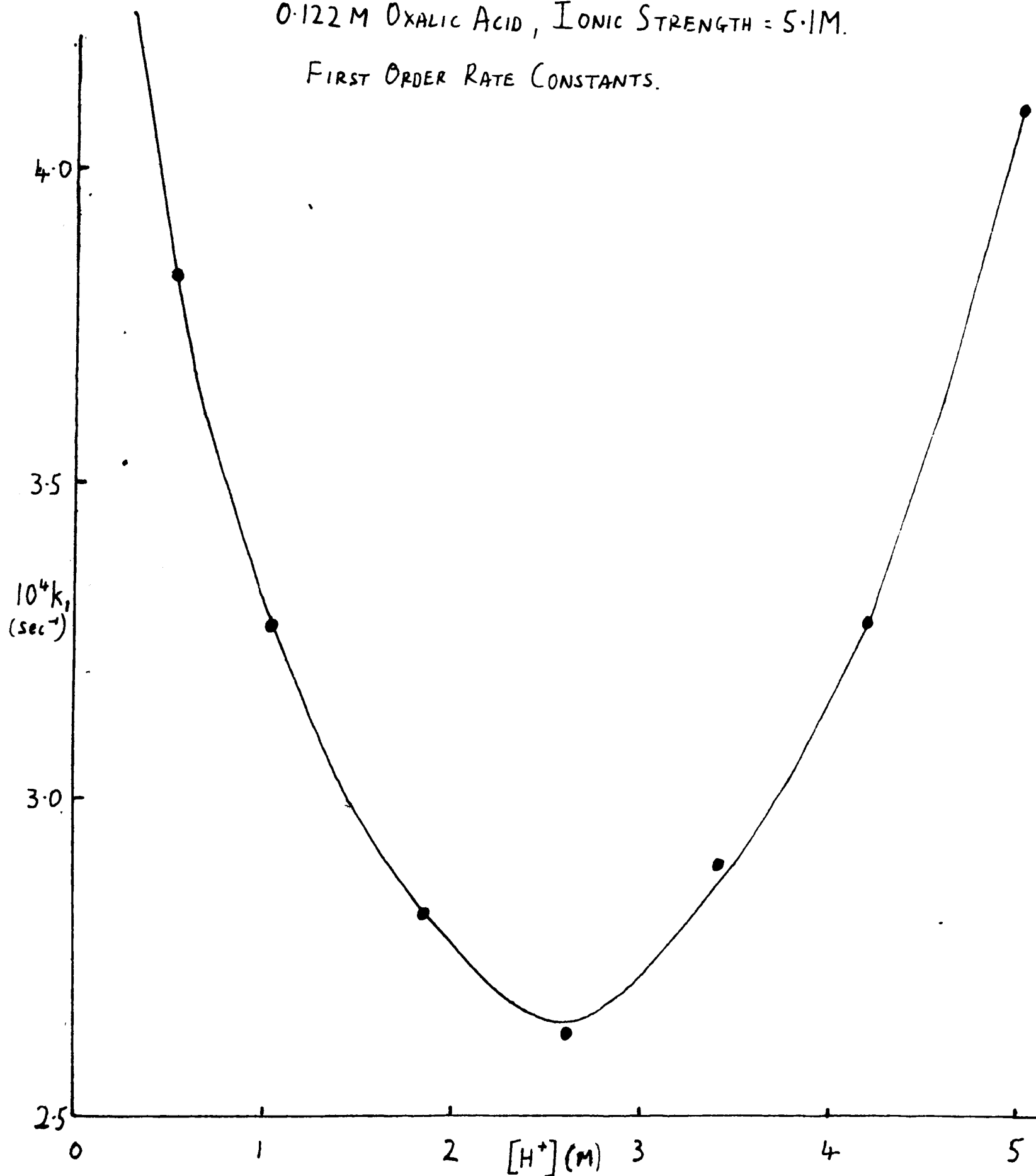
GRAPH 2

OXIDATION OF OXALIC ACID IN AQUEOUS PERCHLORIC ACID.

EFFECT OF VARYING ACID CONCENTRATION AT 50°C.

0.122 M OXALIC ACID, IONIC STRENGTH = 5.1 M.

FIRST ORDER RATE CONSTANTS.

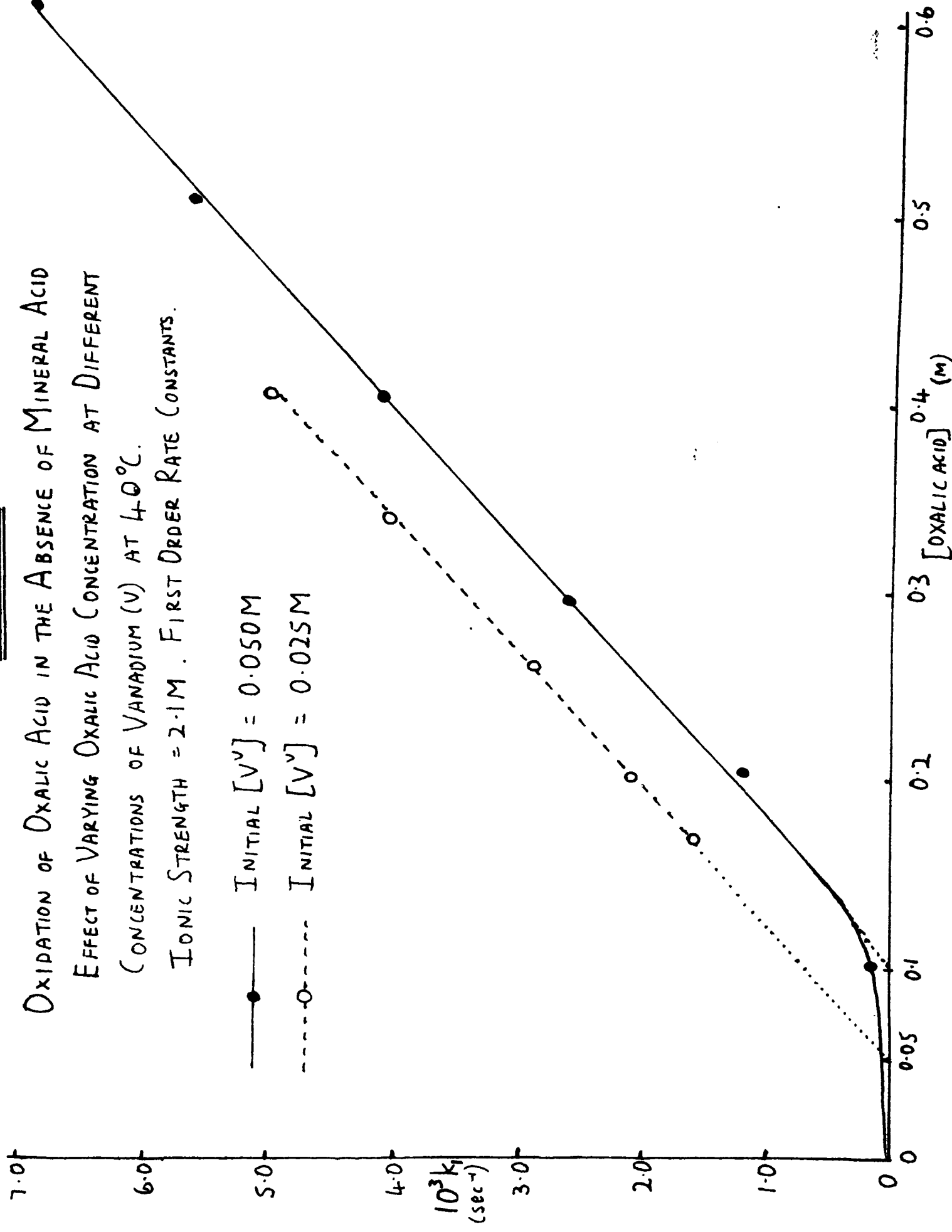


GRAPH 3

OXIDATION OF OXALIC ACID IN THE ABSENCE OF MINERAL ACID
EFFECT OF VARYING OXALIC ACID CONCENTRATION AT DIFFERENT

CONCENTRATIONS OF VANADIUM (V) AT 40°C.

IONIC STRENGTH = 2.1M. FIRST ORDER RATE CONSTANTS.

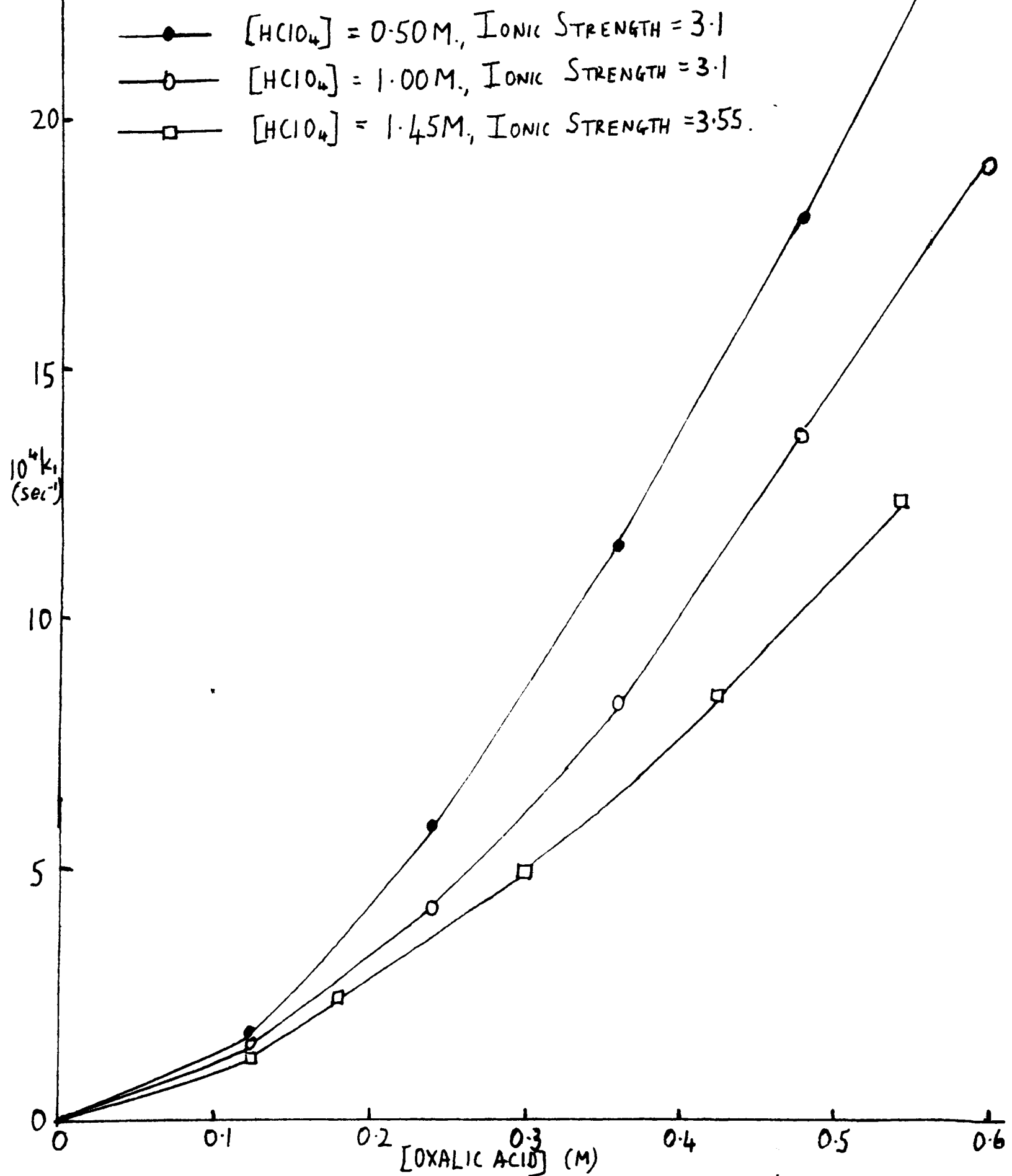


GRAPH 4

OXIDATION OF OXALIC ACID IN DILUTE PERCHLORIC ACID AT 40°C .

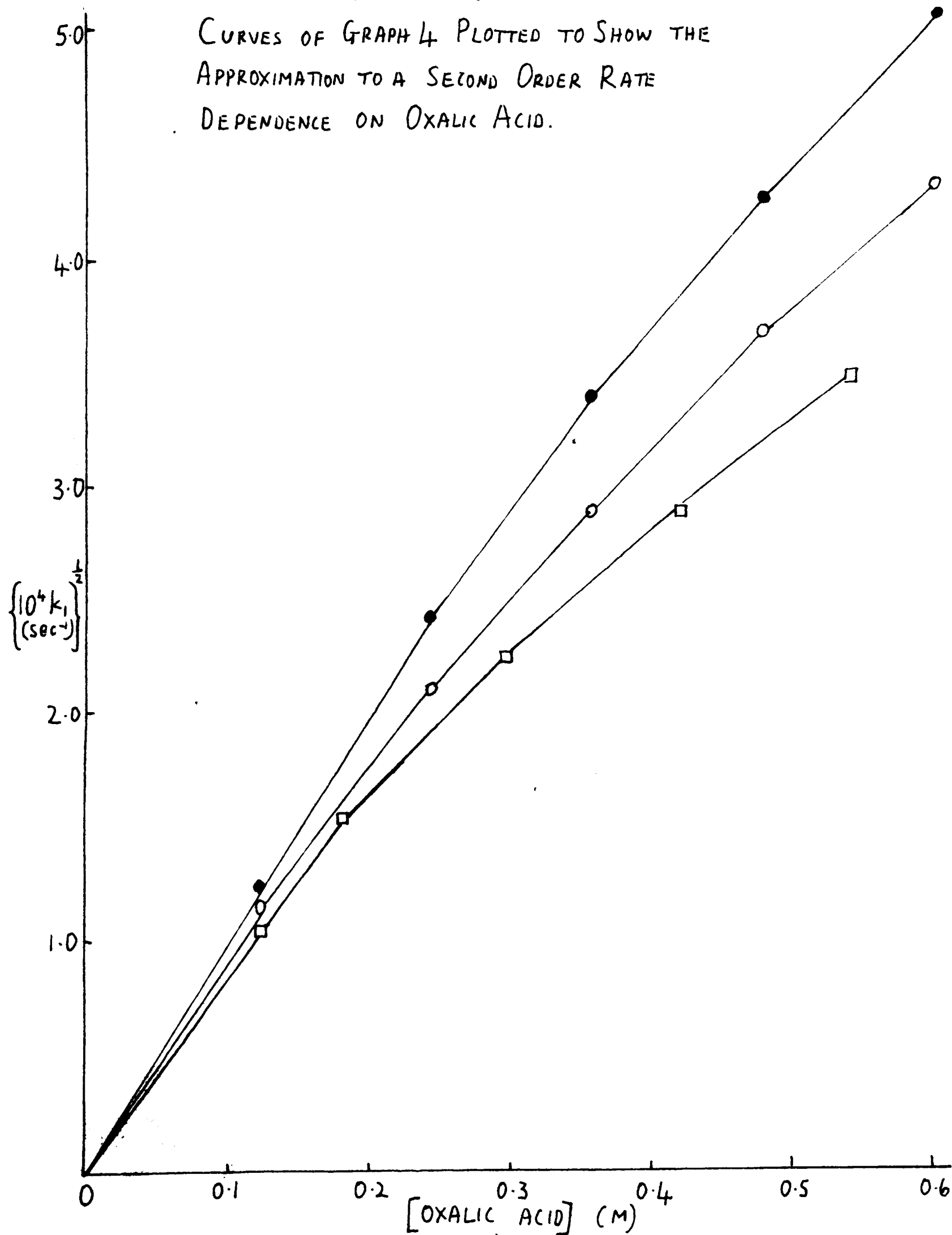
DEPENDENCE OF RATE ON OXALIC ACID CONCENTRATION

INITIAL $[\text{V}^{\text{V}}] = 0.050\text{M}$.



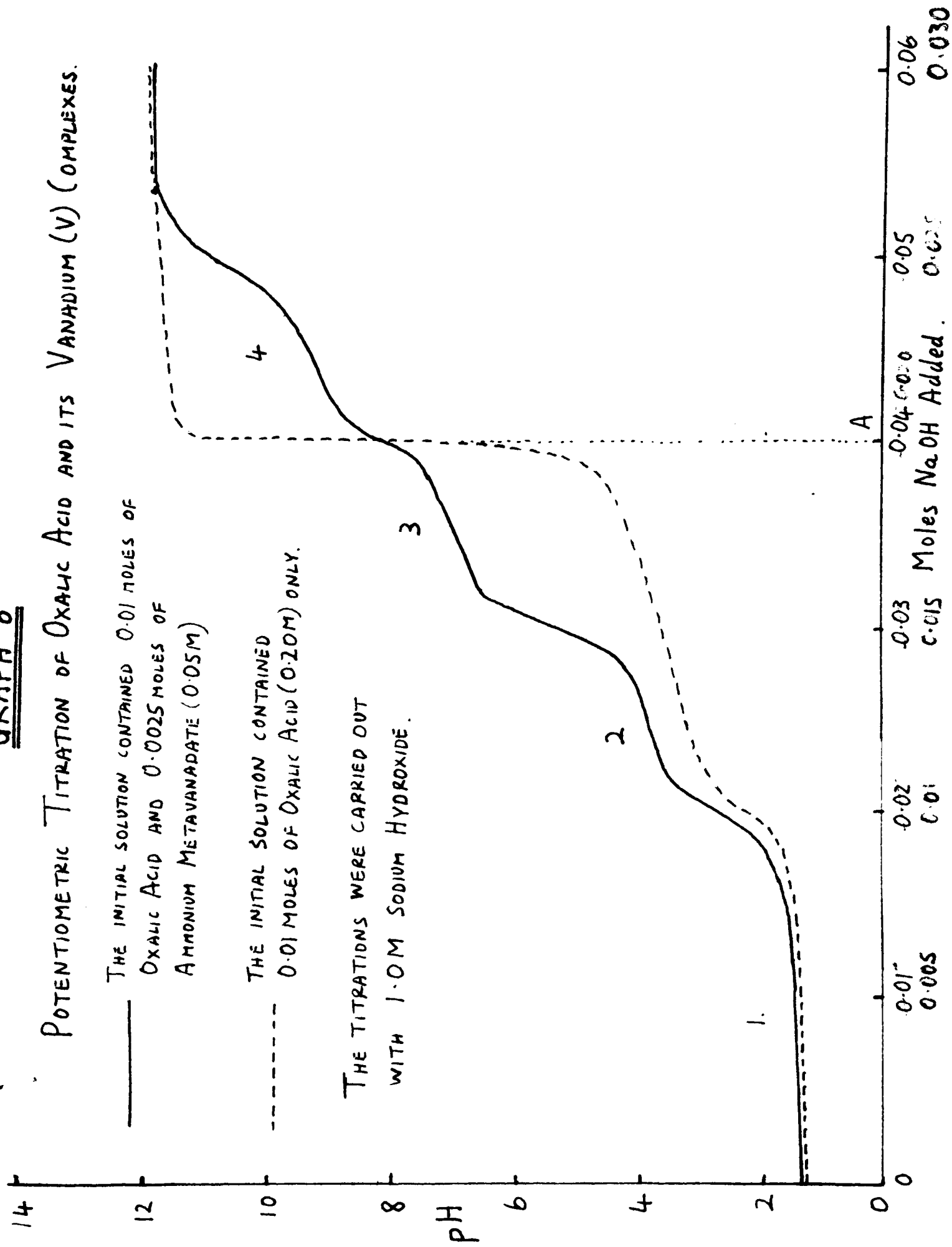
GRAPH 5

CURVES OF GRAPH 4 PLOTTED TO SHOW THE
APPROXIMATION TO A SECOND ORDER RATE
DEPENDENCE ON OXALIC ACID.



GRAPH 6

POTENTIOMETRIC TITRATION OF OXALIC ACID AND ITS VANADIUM (V) COMPLEXES.



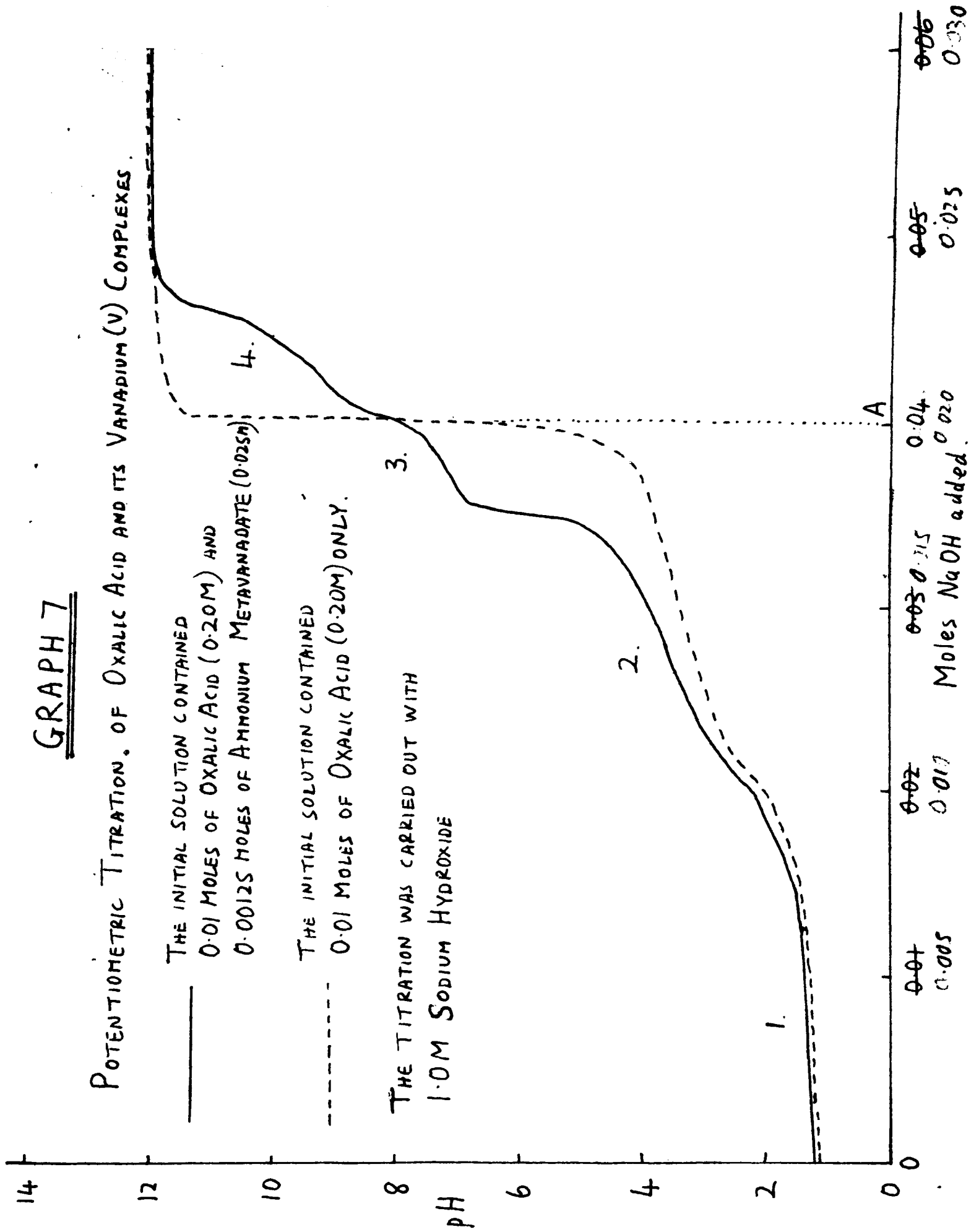
GRAPH 7

POTENTIOMETRIC TITRATION, OF OXALIC ACID AND ITS VANADIUM (V) COMPLEXES.

— THE INITIAL SOLUTION CONTAINED
0.01 MOLES OF OXALIC ACID (0.20M) AND
0.00125 MOLES OF AMMONIUM METAVANADATE (0.025M)

- - - THE INITIAL SOLUTION CONTAINED
0.01 MOLES OF OXALIC ACID (0.20M) ONLY.

THE TITRATION WAS CARRIED OUT WITH
1.0M SODIUM HYDROXIDE



GRAPH 8

OXIDATION OF ACETOIN. DEPENDENCE OF RATE ON

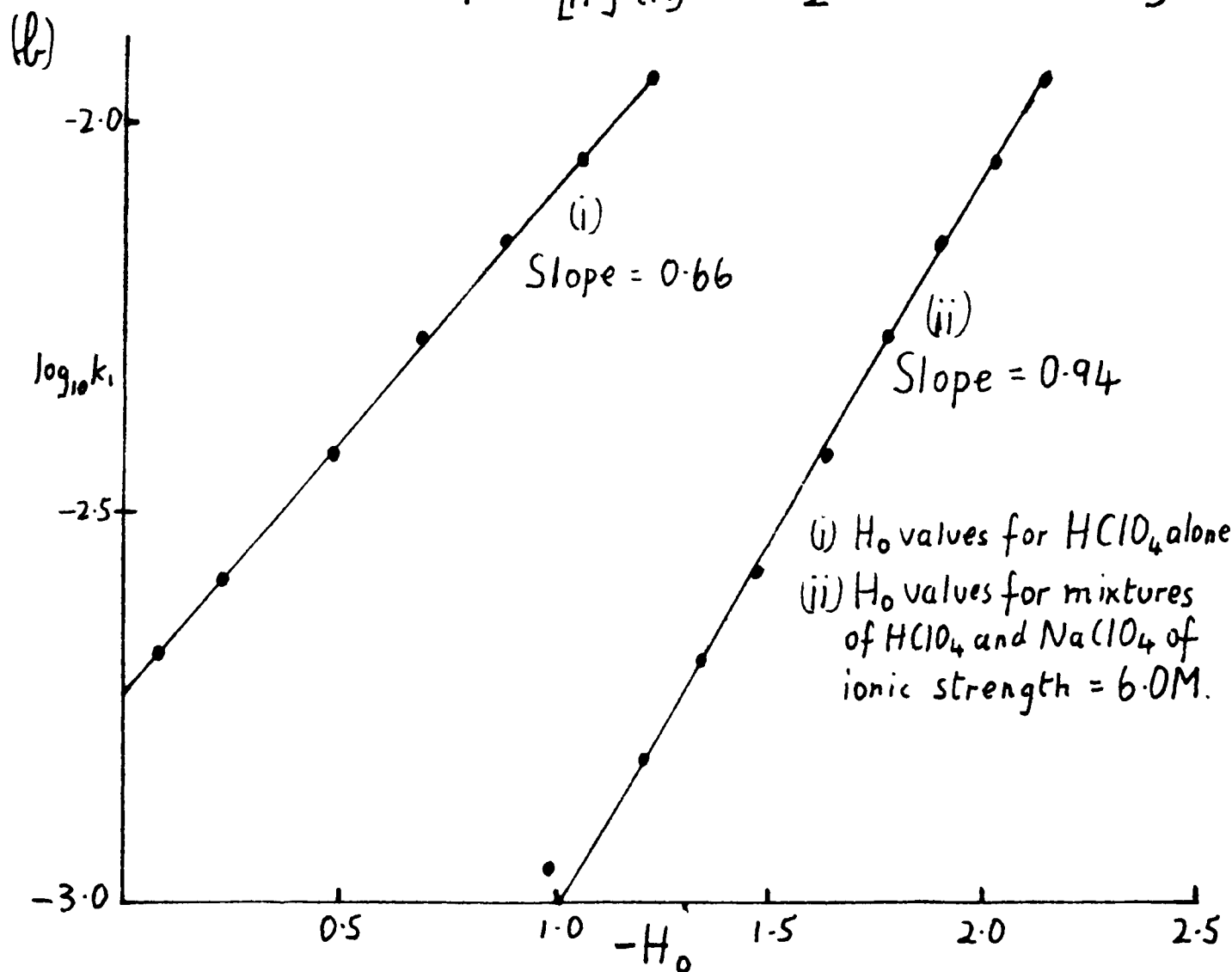
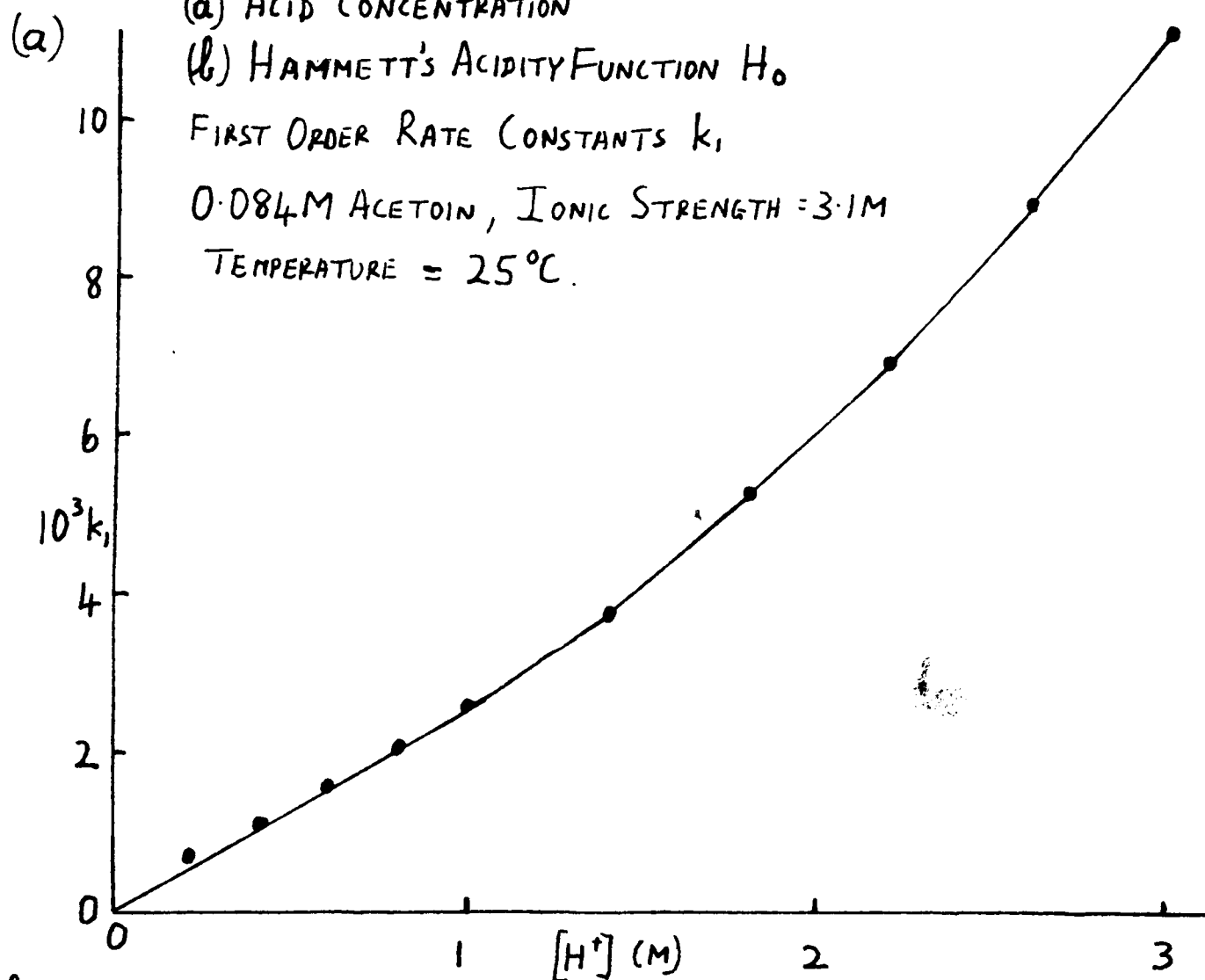
(a) ACID CONCENTRATION

(b) HAMMETT'S ACIDITY FUNCTION H_0

FIRST ORDER RATE CONSTANTS k_1

0.084M ACETOIN, IONIC STRENGTH = 3.1M

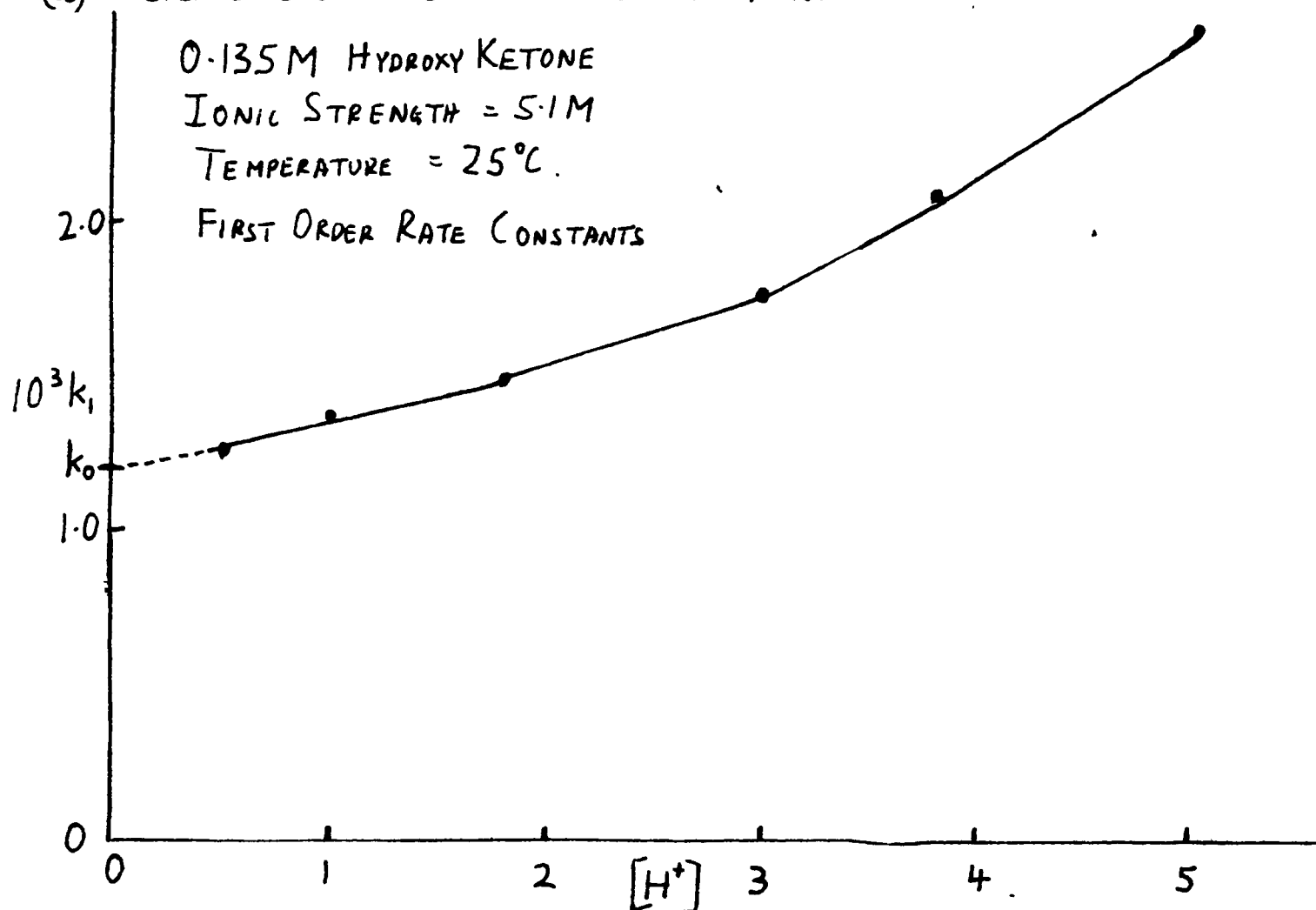
TEMPERATURE = 25°C.



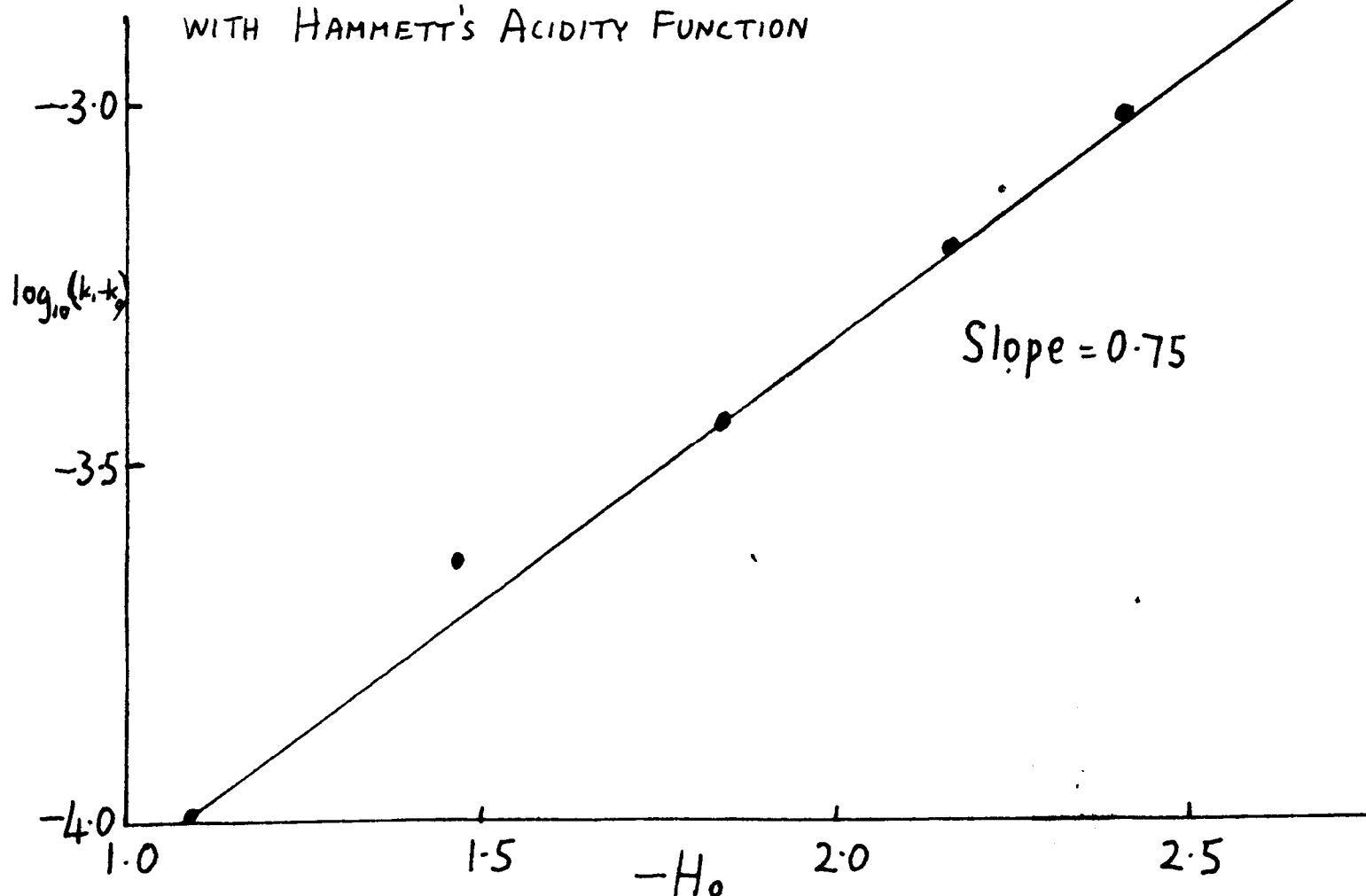
GRAPH 9

OXIDATION OF 3-METHYL 3-HYDROXY 2-BUTANONE IN PERCHLORIC ACID.

(a) DEPENDENCE OF RATE ON ACID CONCENTRATION

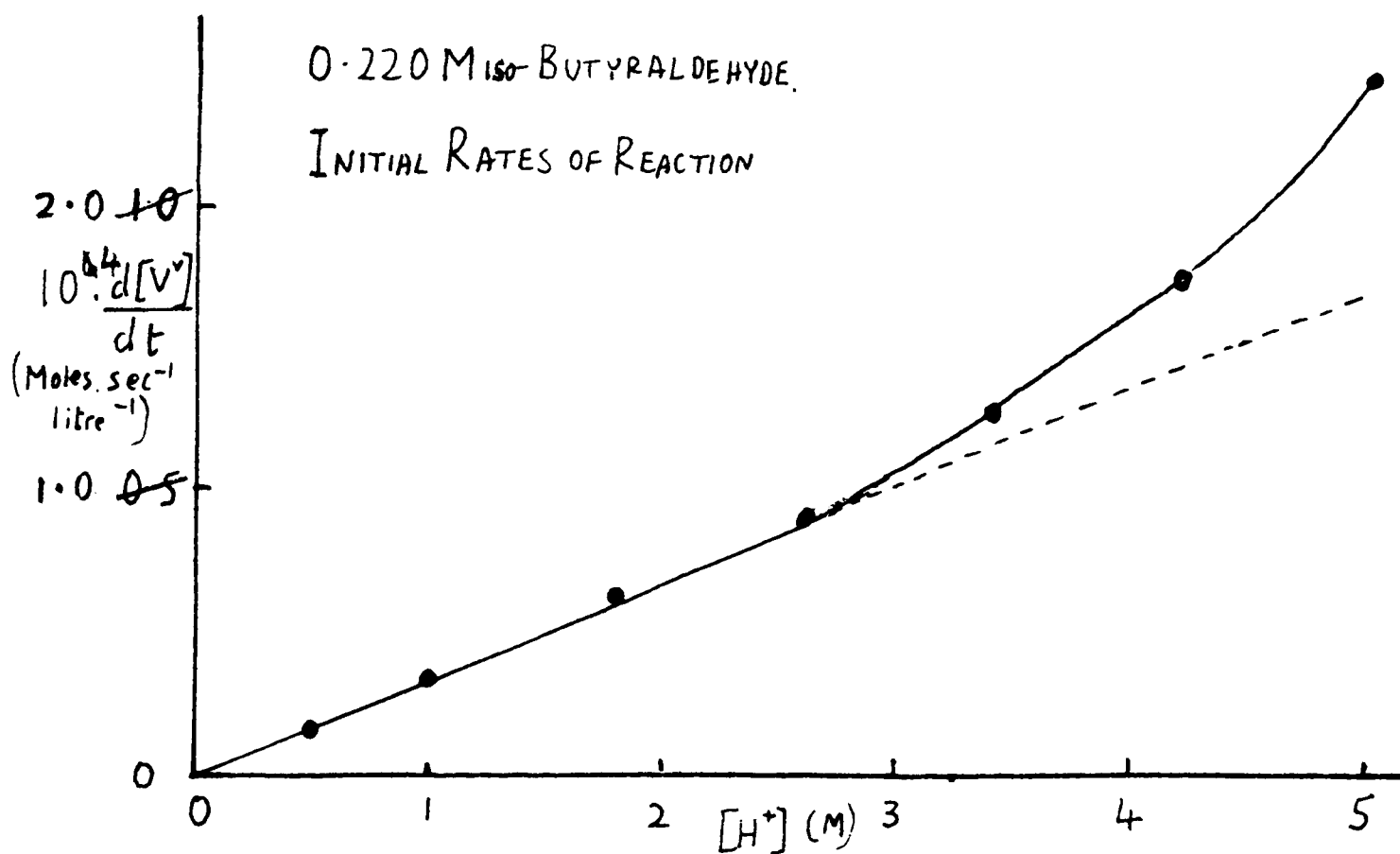
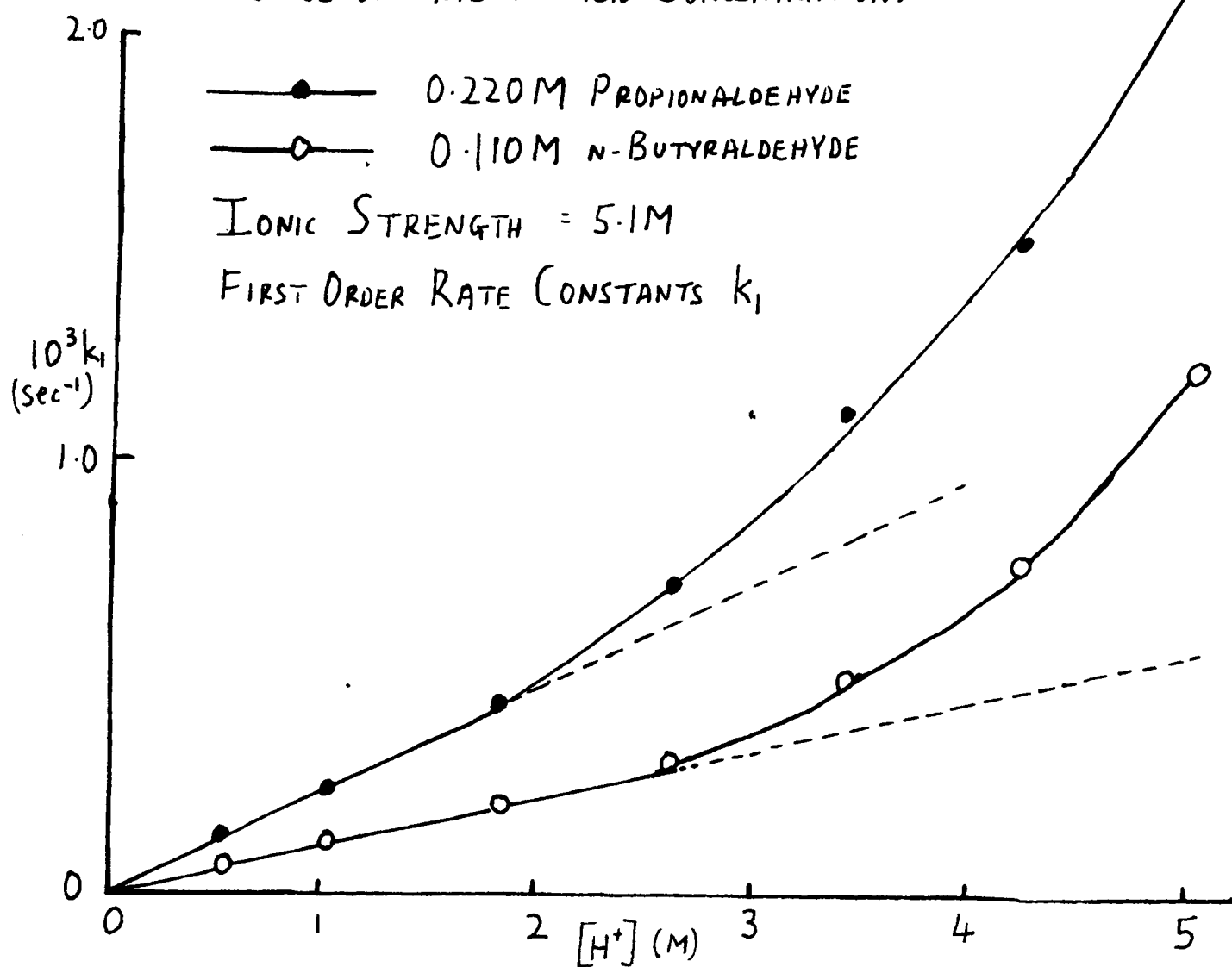


(b) VARIATION OF RATE OF ACID DEPENDENT REACTION
WITH HAMMETT'S ACIDITY FUNCTION



GRAPH 10

OXIDATION OF ALDEHYDES IN PERCHLORIC ACID AT 25°C.
DEPENDENCE OF RATE ON ACID CONCENTRATION.



GRAPH II

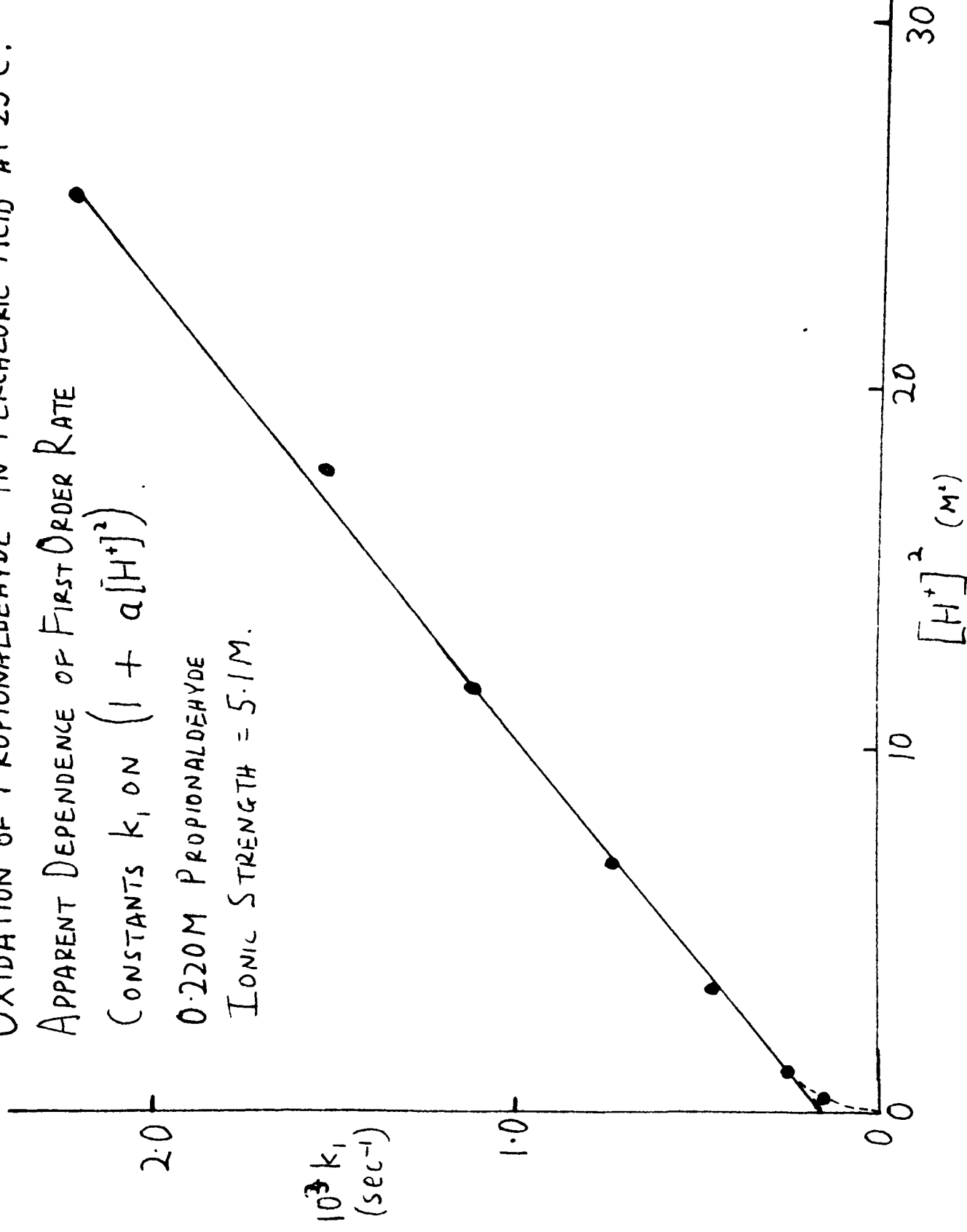
OXIDATION OF PROPIONALDEHYDE IN PERCHLORIC ACID AT 25°C.

APPARENT DEPENDENCE OF FIRST ORDER RATE

CONSTANTS k_1 ON $(1 + a[H^+]^2)$

0.220M PROPIONALDEHYDE

IONIC STRENGTH = 5.1M.



GRAPH 12

REACTION OF ISO-BUTYRALDEHYDE WITH (a) BROMINE, (b) IODINE, (c) VANADIUM(V).

DEPENDENCE OF REACTION RATE ON ACIDITY AT 25°C.

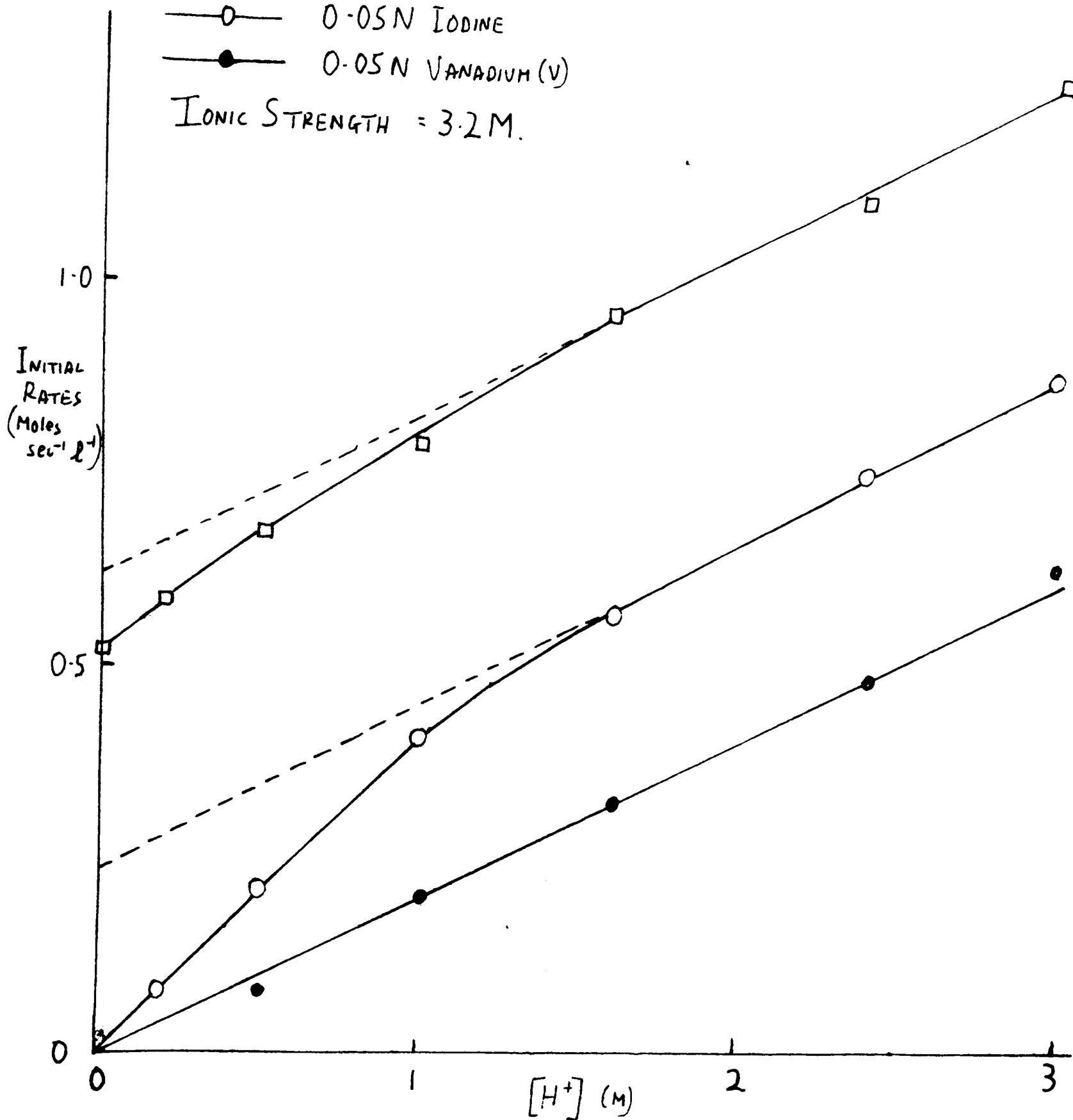
0.220M ISO-BUTYRALDEHYDE

—□— 0.05N BROMINE

—○— 0.05N IODINE

—●— 0.05N VANADIUM(V)

IONIC STRENGTH = 3.2M.



GRAPH 13

ESTIMATION OF ENOLISATION RATES OF ALDEHYDES FROM RATE OF DISAPPEARANCE OF BROMINE (SEE P. 125)

0.110M N-BUTYRALDEHYDE

1.0M PERCHLORIC ACID

0.05N BROMINE

IONIC STRENGTH = 3.1M

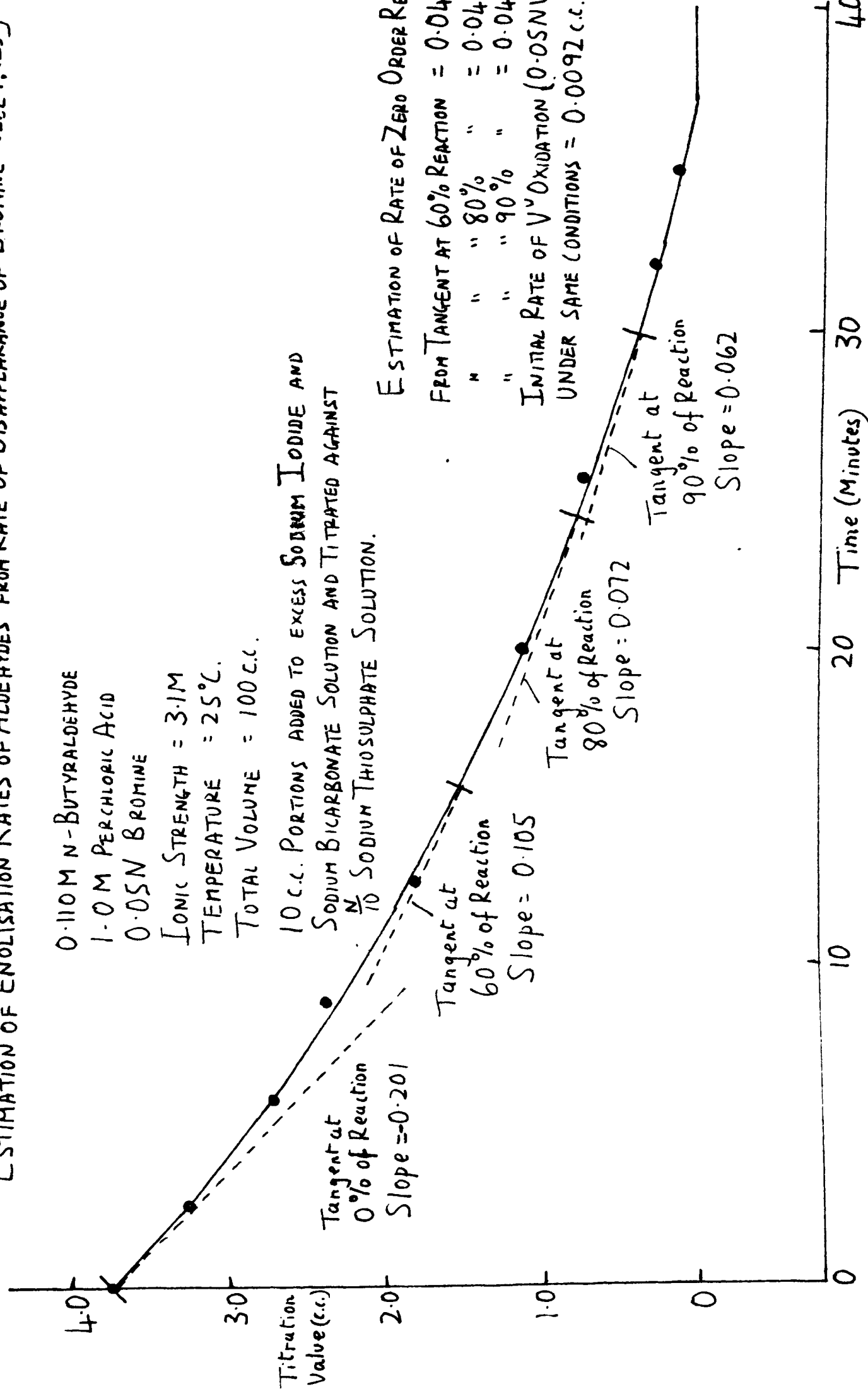
TEMPERATURE = 25°C.

TOTAL VOLUME = 100 c.c.

10 c.c. PORTIONS ADDED TO EXCESS SODIUM IODIDE AND

SODIUM BICARBONATE SOLUTION AND TITRATED AGAINST

$\frac{N}{10}$ SODIUM THIOSULPHATE SOLUTION.



REFERENCES

1. Berzelius, *Traite de Chemie* Vol. IV, p.642 (Paris 1833).
2. J.S.Littler, B.Sc. Thesis (Oxford 1959).
3. Bobtelsky and Glasner, *J.A.C.S.*, 1942, 64, 1462.
4. Littler and Waters, *J.C.S.*, 1959, 1299, 3014, 4046.
5. Littler, Mallet and Waters, *J.C.S.*, 1960, 2761;
Littler and Waters, *J.C.S.*, 1960, 2767.
6. Ducret, *Ann. Chim. (France)* 1951, 6, 705.
7. Bullberg, *Z. Phys. Chem.* 1903, 45, 124;
Meyer and Pawletta, *Z. Phys. Chem.* 1927, 125, 49.
8. Rossotti and Rossotti, *Acta. Chem. Scand.*, 1956, 10, 957.
9. Foerster and Botcher, *Z. Phys. Chem.*, 1930, 151, 321;
Corzelli and Yost, *J.A.C.S.*, 1933, 55, 1909;
Hart and Partington, *J.C.S.*, 1940, 1532.
Carpenter, *J.A.C.S.*, 1934, 56, 1349.
Britton and Welford, *J.C.S.* 1940, 895.
10. La Salle and Cobble, *J. Phys. Chem.*, 1955, 59, 519
11. J.S.Littler, Private Communication.
12. West and Skoog, *Analyt. Chem.*, 1959, 31, 583;
West and Skoog, *J.A.C.S.*, 1960, 82, 280.
13. See e.g. R.P.Bell, "Acid-Base Catalysis" (Oxford 1941)
Ch. 2.
14. Kirk and Browne, *J.A.C.S.*, 1928, 50, 337;

- Higginson, Sutton and Wright, J.C.S., 1953, 1380,1402;
Higginson and Marshall, J.C.S., 1957, 477;
Saffir and Taube, J.A.C.S., 1960, 82, 13.
15. Drummond and Waters, J.C.S., 1953, 2836.
 16. Merz and Waters, J.C.S., 1949, S 15.
 17. Mackinnon and Waters, J.C.S., 1953, 323.
 18. Drummond and Waters, J.C.S., 1955, 497.
 19. Waters, Trans. Farad. Soc., 1946, 42, 184.
 20. Drummond and Waters, J.C.S., 1953, 435.
 21. Bigeleisen, J. Chem. Phys., 1949, 17, 675.
 22. Bell, Fendley and Hulett, Proc. Roy. Soc., 1956,
A 235, 453.
 23. Wiberg, Chem. Rev., 1955, 55, 713.
 24. Halevi, Tetrahedron, 1957, 1, 174.
 25. Weston, Tetrahedron, 1959, 6, 31.
 26. Shiner, J.A.C.S., 1953, 75, 2925;
Bartell, Tet. Letters, 1960, 6, 13.
 27. Seltzer, Chem. and Ind., 1959, 1313.
 28. Smith, Bowman and Kinet, J.A.C.S., 1959, 81, 997.
 29. Moelwyn-Hughes and Bonhoeffer, Z. Electrochem., 1934,
40, 469.
 30. Thomas and Turner, Q. Rev., 1953, 7, 407.
 31. Bunton and Shiner, J.C.S., 1960, 1596.
 32. Wiberg and Saegbarth, J.A.C.S., 1957, 79, 2822.
 33. R.P.Bell, "The Proton in Chemistry", (Methuen 1960)

34. Paul and Long, Chem. Rev., 1957, 57, 1, 935.
35. Hammett, "Physical Organic Chemistry", (McGraw Hill, New York, 1940) Ch. 9;
Hammett and Deyrup, J.A.C.S., 1932, 54, 2721;
Hammett, Chem. Rev., 1935, 16, 67.
36. Hammett and Paul, J.A.C.S., 1934, 56, 827;
Bascombe and Bell, J.C.S., 1959, 1096.
37. Zucker and Hammett, J.A.C.S., 1939, 61, 2791.
38. Taft, Deno and Skell, Ann. Rev. Phys. Chem., 1958, 9, 303.
39. Whalley, Trans. Farad. Soc., 1959, 55, 815;
Bunton and Llewellyn, J.C.S., 1957, 3402.
40. Westheimer et al., J.A.C.S., 1949, 71, 25; 1951, 73, 6;
1952, 74, 4387; 1958, 80, 3030.
41. Rocek and Kupricka, Coll. Czech. Chem. Comm., 1958, 23, 2068;
Chem. Listy, 1958, 52, 1735.
42. Venkatasabramanian, Proc. Ind. Acad. Sci., 1959, A 50, 156.
43. Kwart and Francis, J. A.C.S., 1959, 81, 2116.
44. Brownell, Leo, Chang and Westheimer, J.A.C.S., 1960, 82, 406.
45. Westheimer, Chem. Rev., 1949, 45, 419.

46. Waters, Q. Rev., 1958, 12, 277.
47. Chang and Westheimer, J.A.C.S., 1960, 82, 1401.
48. Chatterji and Mukherjee, Z. Phys. Chem. (Leipzig), 1958, 208, 281.
49. Wiberg and Mill, J.A.C.S., 1958, 80, 3022.
50. Graham and Westheimer, J.A.C.S., 1958, 80, 3030.
51. Rodek, Tet. Letters, 1959, 5, 1.
52. Wiberg and Evans, Tetrahedron, 1960, 8, 313.
53. Drummond and Waters, J.C.S., 1953, 3119.
54. Drummond and Waters, J.C.S., 1954, 2456.
55. Levesley and Waters, J.C.S., 1955, 217.
56. Noyes, Trans. New York Acad. Sci., 1951, 2, 13, 314;
Malcolm and Noyes, J.A.C.S., 1952, 74, 2769;
Ladbury and Cullis, Chem. Rev., 1958, 58, 403.
57. Taube, J.A.C.S., 1948, 70, 1216.
58. Wiberg and Stewart, J.A.C.S., 1955, 77, 1786.
59. Higginson, Sutton and Wright, J.C.S., 1953, 1380, 1402;
Duke and Forrist, J.A.C.S., 1949, 71, 2790;
Hargreaves and Sutcliffe, Trans. Farad. Soc., 1955, 51,
786, 1105;
Ardon, J.C.S., 1957, 1811;
Venkatakrisnan and Santappa, Z. Phys. Chem. (Frankfurt),
1958, 16, 73.

60. Shorter and Hinshelwood, J.C.S., 1950, 3273;
Shorter, J.C.S., 1950, 3425.
61. Speakmann and Waters, J.C.S., 1955, 40;
Conant and Aston, J.A.C.S., 1923, 50, 2783.
62. Cordner and Pausacker, J.C.S., 1953, 102;
Angyal and Young, J.A.C.S., 1959, 81, 5251, 5467.
63. Bell, Rivlin and Waters, J.C.S., 1958, 1696.
64. Duke, J. A.C.S., 1947, 69, 3054;
Duke and Bulgrin, J.A.C.S., 1954, 76, 3803;
Buist and Bunton, J.C.S., 1954, 1406;
Buist, Bunton and Miles, J.C.S., 1957, 4567, 4575.
65. Rigby, J. C.S., 1950, 1907.
66. Thomas, Trudel and Bywater, J. Phys. Chem., 1960, 64,
51.
67. Marshall and Waters, J.C.S., 1960, 2392.
68. Mosher and Whitmore, J.A.C.S., 1948, 70, 2544;
Mosher and Langerak, J.A.C.S., 1949, 71, 286; 1951, 73,
1302.
69. Hampton, Leo and Westheimer, J.A.C.S., 1956, 78, 306.
70. Celeste, Miss. Abs., 1960, 20, 2551.
71. Neidig, Funck, Uhrich, Baker and Kreiser, J.A.C.S.,
1950, 72, 4617.
72. Ford, Thompson and Marvel, J.A.C.S., 1935, 57, 2621.
73. Norris and Rigby, J.A.C.S., 1932, 54, 2098.

74. Skell and Hauser, J.A.C.S., 1942, 64, 2633;
Whitmore and Rothrock, J.A.C.S., 1932, 54, 3431;
Whitmore, Wittle and Popkin, J.A.C.S., 1939, 61, 1586;
Dostrovsky and Samuel, J.C.S., 1954, 658.
75. Hammett and Paul, J.A.C.S., 1934, 56, 830.
76. Chapman and Smith, J.C.S., 1867, 20, 173.
77. Jones, Waters and Littler, J.C.S., 1961, 630.
78. Welch and Clemo, J.C.S., 1928, 2629.
79. Rule and Harrower, J.C.S., 1930, 2325.
80. Dhar, J.C.S., 1917, 707.
81. Morette and Gaudefroy, Bull. Soc. Chim., 1954, 956.
82. Horne, J. Phys. Chem., 1960, 64, 1512.
83. Weiss, Disc. Farad. Soc., 1947, 2, 188.
84. Barr, J.A.C.S., 1940, 62, 1604.
85. Pound, J. Phys. and Colloid Chem., 1947, 51, 1449.
86. Rigby, J.C.S., 1951, 793.
87. Dirscherl and Schöllig, Ber., 1938, 71B, 418.
88. Bailar, "Chemistry of the Coordination Compounds"
(New York 1956).
89. "Interatomic Distances", Chem. Soc. Special Publication
No. 11, 1958.
90. J.S.Littler, Private Communication.
91. Campbell, Campbell and Eby, J.A.C.S., 1938, 60, 2882.

92. Froning and Hennion, J.A.C.S., 1940, 62, 653.
93. Linstead, Elvidge and Whalley, "Modern Techniques of Organic Chemistry" (Butterworth, London, 1955) p.104.
94. Ansell, Hickinbottom and Hyatt, J.C.S., 1955, 1593.
95. Bawn and White, J.C.S., 1951, 331, 339, 343;
Bawn and Williamson, Trans. Farad. Soc., 1951, 47, 721.
96. Ardon, J.C.S., 1957, 1811.
97. Littler, J.C.S., 1959, 4135.
98. Herold, Z. Phys. Chem., B, 1932, 18, 265
99. Bell and Clunie, Trans. Farad. Soc., 1952, 48, 439.

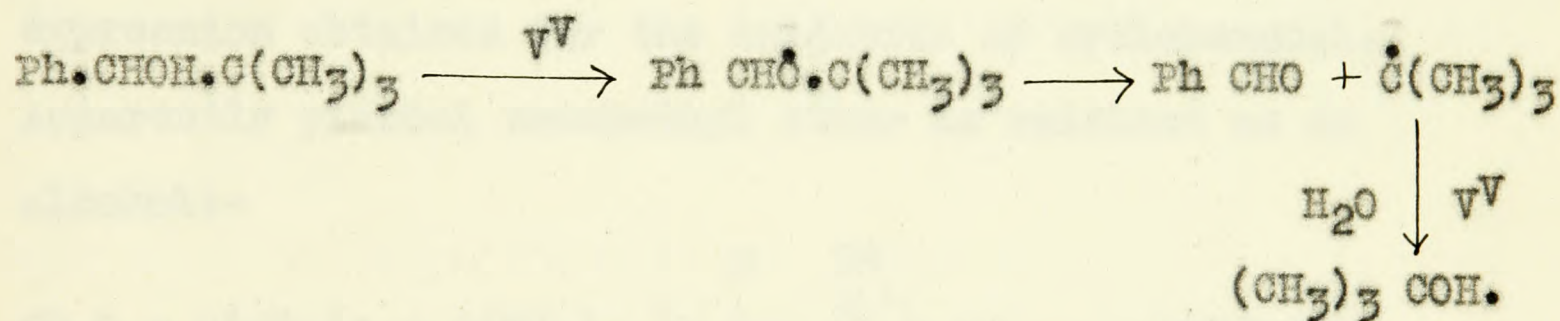
SUMMARY

Vanadium (V) only possesses oxidising properties in acid solutions. The main species present in these conditions is the cation VO_2^+ , but this is protonated to a small extent to give the more powerfully oxidising species $\text{V}(\text{OH})_2^{2+}$.

Previous work in this laboratory has shown that vanadium (V) is a one-electron oxidising agent and is reduced by organic compounds to vanadium (IV) only. No vanadium (III) has yet been detected in any of these oxidation reactions while the presence of free radicals, which are usually associated with one-electron oxidations, has been demonstrated in many vanadium (V) oxidations of organic compounds.

To extend the knowledge of these oxidations a number of alcohols, α -hydroxy acids, acyloins and aldehydes, and also pinacol and oxalic acid have been oxidised by acid solutions of vanadium (V). The products of these oxidations have been investigated and in many cases a detailed investigation of the kinetics of the reactions has been made.

1. α -t-Butylbenzyl alcohol $\text{PhCH(OH).C(CH}_3)_3$ was shown to be oxidised initially to benzaldehyde with direct carbon-carbon bond fission. The yield of pivalophenone $\text{Ph CO.C(CH}_3)_3$, which is not oxidised under the conditions used, was only about 10-15%. This direct carbon-carbon bond fission has been noted previously in the oxidation of various secondary alcohols with chromium (VI) and manganese (VII), but in general the yield of fission products is much smaller with these reagents. The following mechanism is proposed for the fission reaction in the vanadium (V) oxidation.



2. 2-Phenylcethyl alcohol was also found to be oxidised by vanadium (V) with carbon-carbon bond fission. Benzaldehyde was the main product of the reaction. Phenylacetic acid, which is not oxidised by vanadium (V) was not found in the reaction products.

3. The oxidations of the α -hydroxy acids lactic acid, malic acid, mandelic acid and α -hydroxy iso-butyric

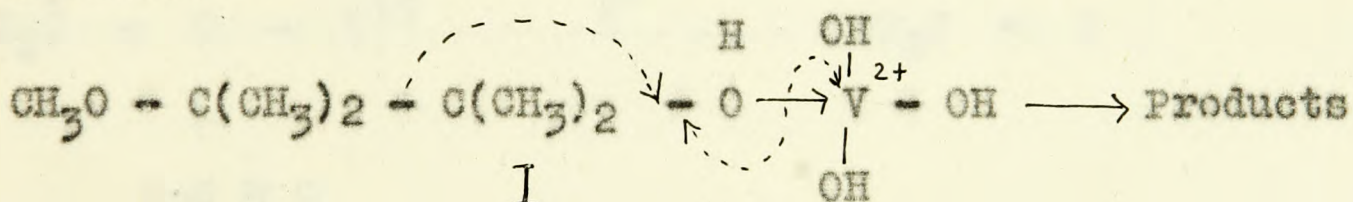
acid, and also of pinacol, were found to follow the rate equation

$$\frac{-d[V^V]}{dt} = k [V^V][\text{Substrate}] (1 + a h_0)$$

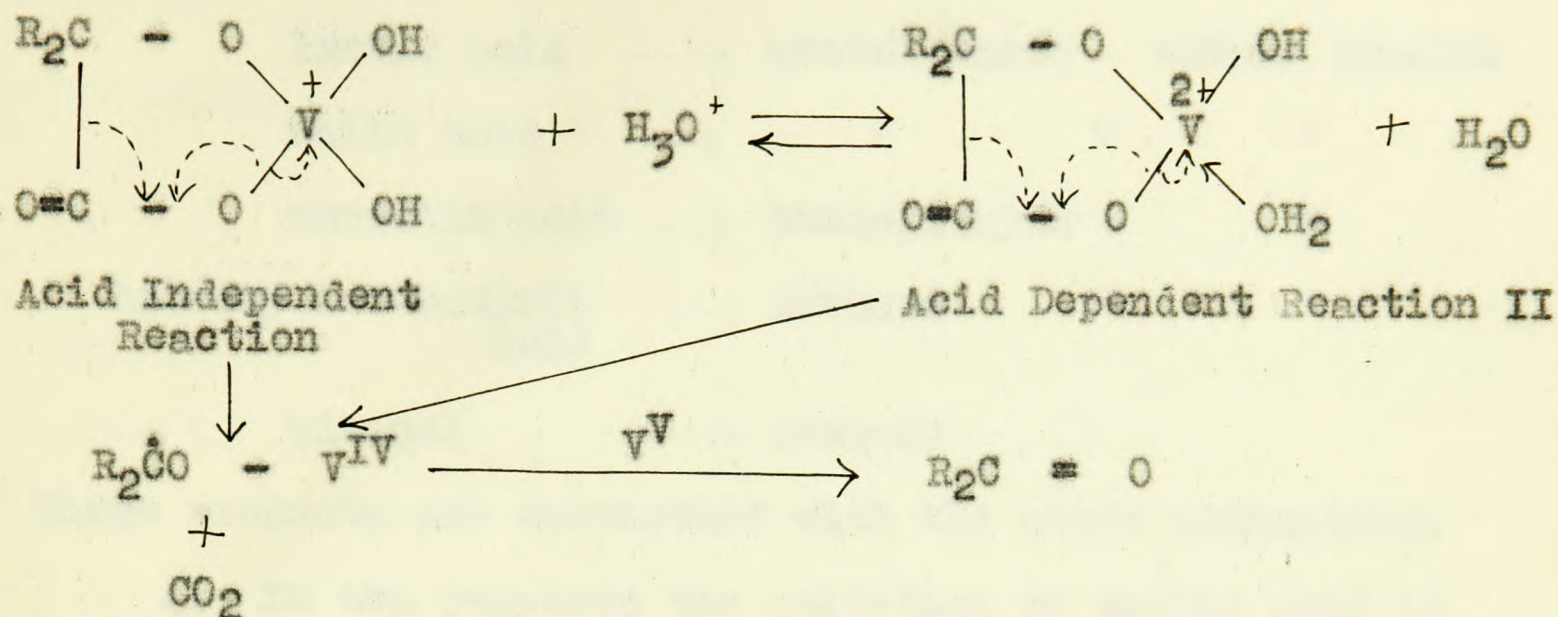
where a is a constant and h_0 is the Hammett acidity function. The acid independent reaction must be due to VO_2^+ while the acid dependent reaction must be due to $V(OH)_2^{2+}$. The above expression contrasts with that obtained for the oxidation of pinacol monomethyl ether:-

$$\frac{-d[V^V]}{dt} = k [V^V][\text{ether}][H^+]$$

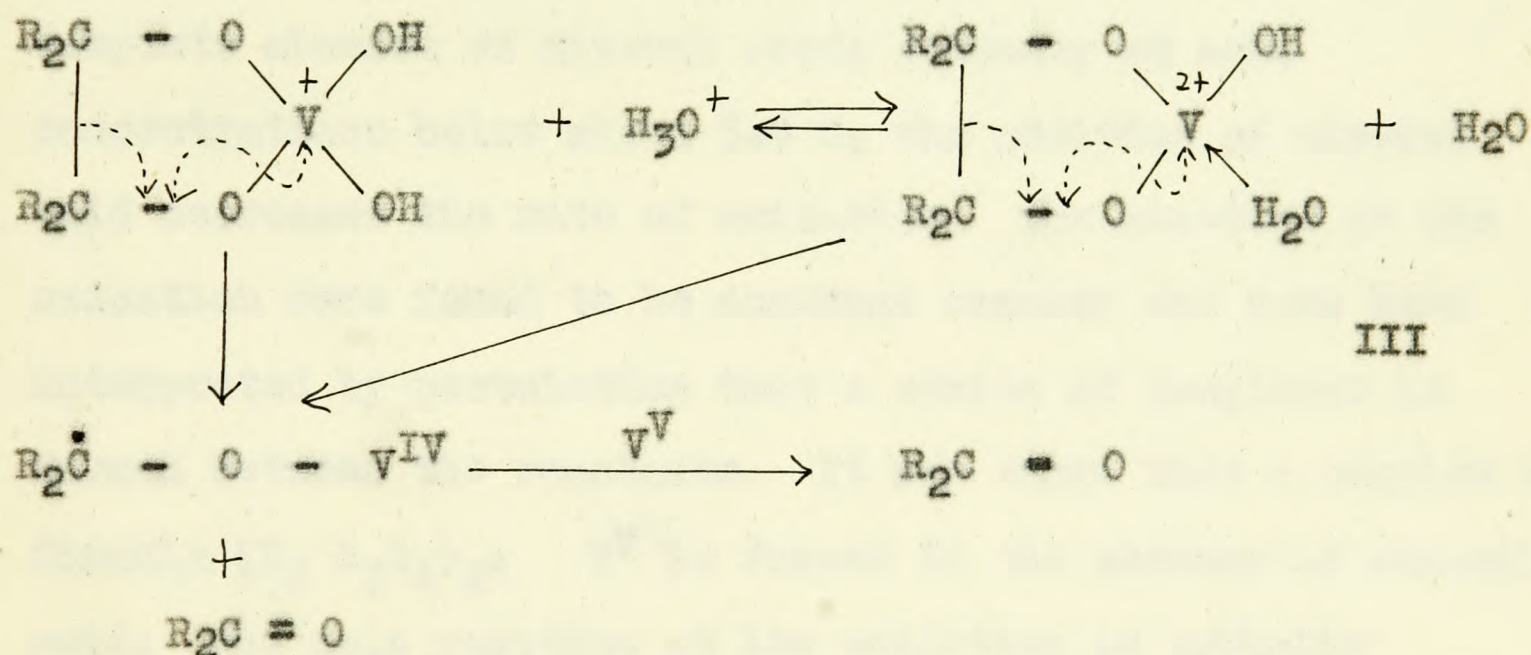
This has no acid independent term and is the same as the expression obtained for the oxidation of cyclohexanol. Apparently pinacol monomethyl ether is oxidised as an alcohol:-



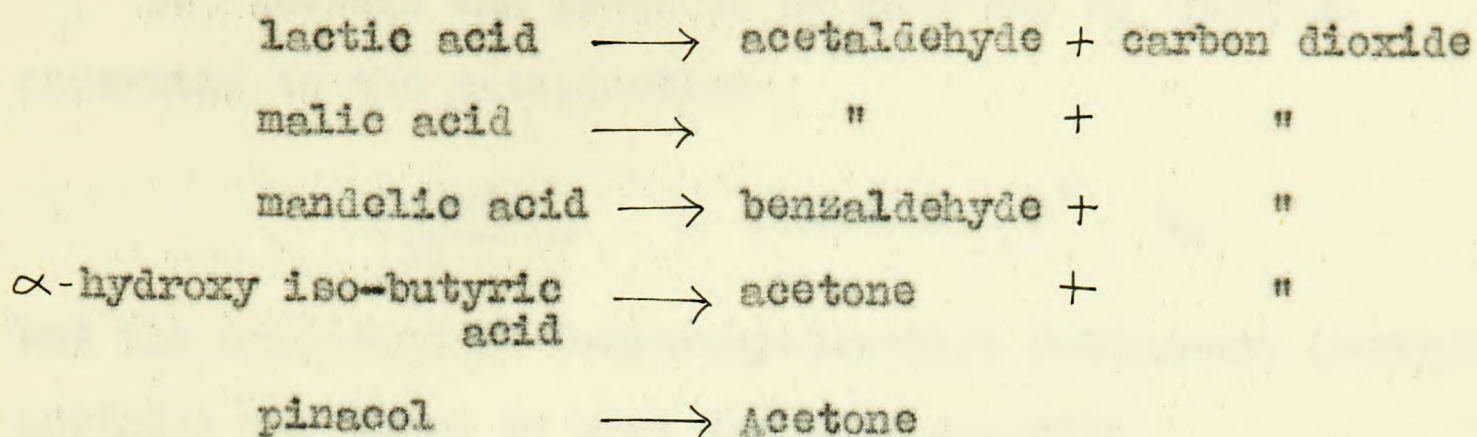
The mechanism of oxidation of the hydroxy acids is postulated to be



and for pinacol

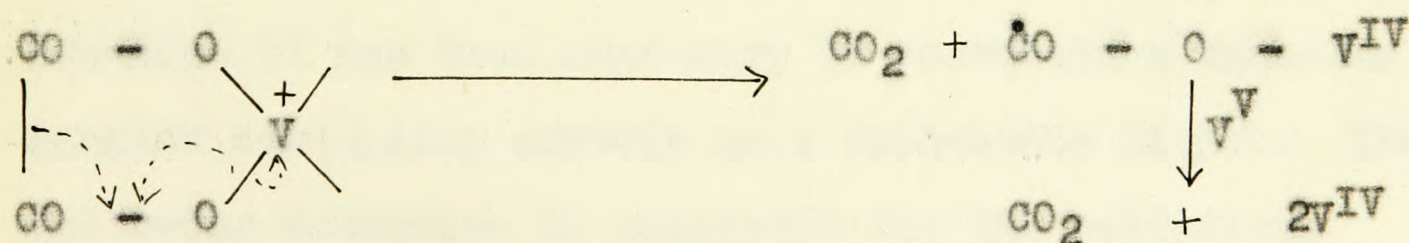


Complex I contains a molecule of water which is not present in II or III and this fact has been used to explain the difference in the observed acidity dependences of the two reactions. The products of the oxidations were shown to be



These products are consistent with the above mechanisms.

4. In two respects the oxidation of oxalic acid is unique among vanadium (V) oxidations which have been investigated so far; firstly oxidation can occur in the complete absence of mineral acid; secondly at acid concentrations below about 3.5 M, the addition of mineral acid decreases the rate of oxidation. The kinetics of the oxidation were found to be somewhat complex and have been interpreted by postulating that a series of complexes is formed between the reactants. It was shown that a complex of formula $(\text{H}_2\text{C}_2\text{O}_4)_2 \cdot \text{V}^{\text{V}}$ is formed in the absence of mineral acid. The main reaction at low acidities is probably assisted by the formation of a cyclic chelate structure



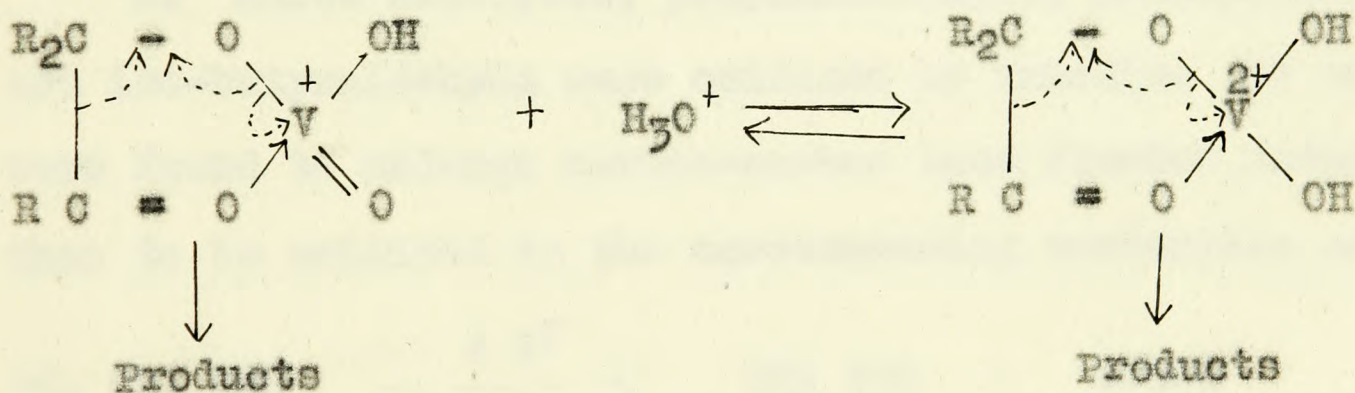
5. Acetoin was found to be oxidised by vanadium according to the rate equation

$$\frac{-d[V^V]}{dt} = k [\text{Acetoin}] [V^V] \cdot h_0$$

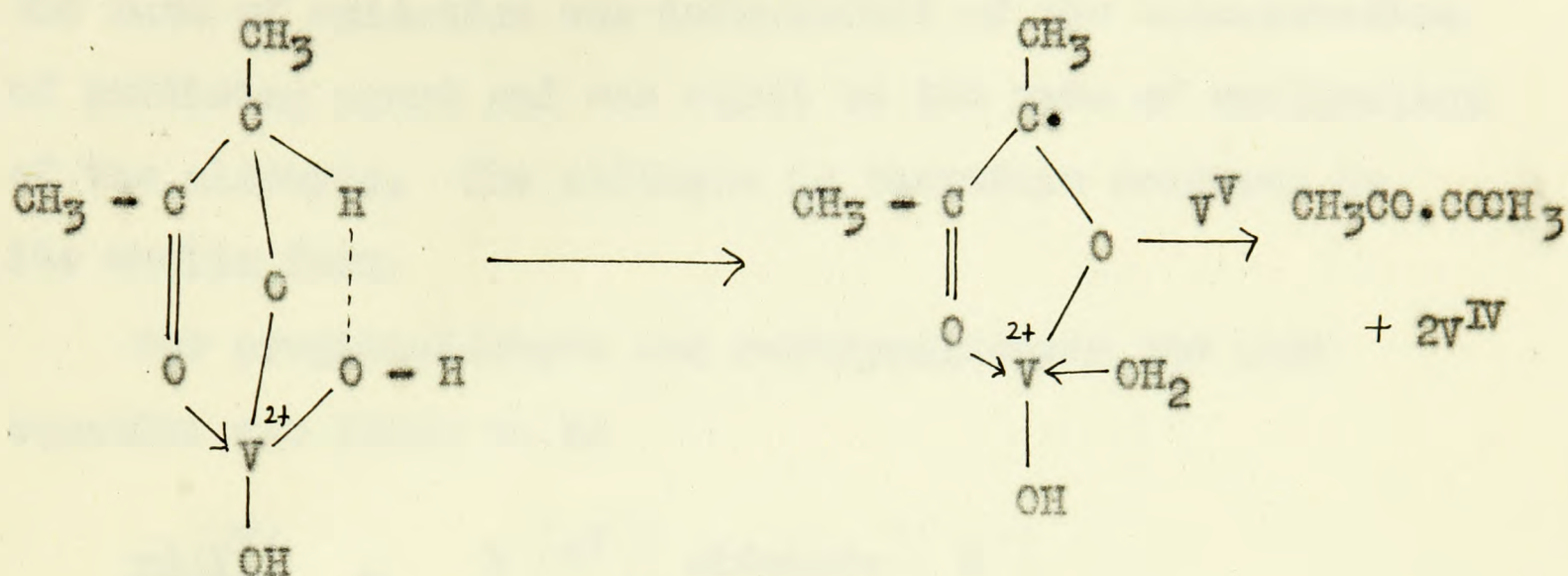
but the oxidation of 3-hydroxy-3-methyl 2-butanone (methyl acetoin) was found to have the rate equation

$$\frac{-d[V^V]}{dt} = k [\text{Methyl acetoin}] [V^V] \cdot (1 + a h_0)$$

Methyl acetoin must therefore be oxidised in a similar manner to the hydroxy acids giving acetone and acetic acid as products.

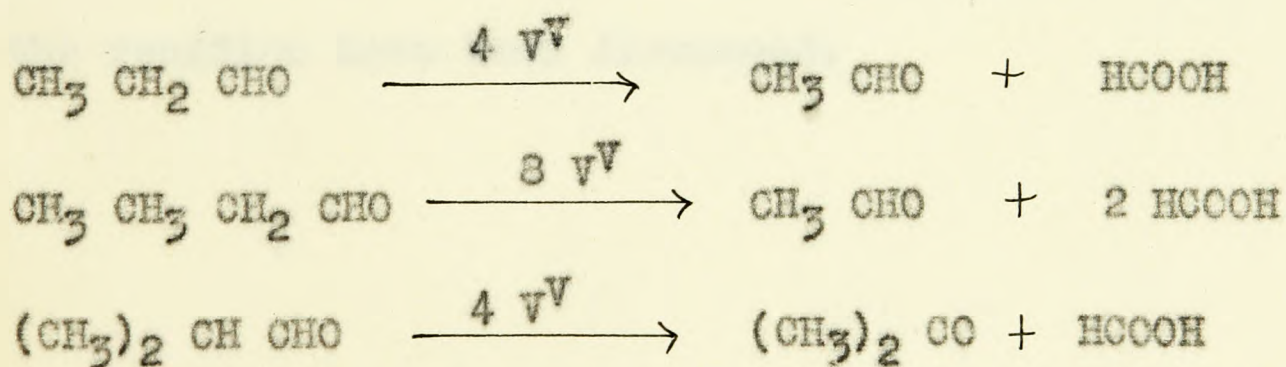


Acetoin was found to be oxidised to diacetyl initially. In order to account accurately for the observed oxidation kinetics it has been necessary to postulate a vanadium (V) complex containing acetoin as a tridentate ligand. The following mechanism is suggested for the oxidation,



and in the discussion it is explained that this mechanism is equivalent to an oxidation of acetoin both as a ketone and as an alcohol.

6. Three aldehydes, propionaldehyde, n-butyraldehyde and iso-butyraldehyde were oxidised by vanadium (V) and were found to undergo carbon-carbon bond fission rather than to be oxidised to the corresponding carboxylic acids.



The kinetics of the iso-butyraldehyde oxidation followed the rate equation

$$-\frac{d[\text{V}^{\text{V}}]}{dt} = k [\text{aldehyde}] [\text{H}^+]$$

The rate of oxidation was independent of the concentration of oxidising agent and was equal to the rate of enolisation of the aldehyde. The aldehyde is therefore oxidised in its enolic form.

For propionaldehyde and n-butyraldehyde the rate equation was found to be

$$\frac{-d[y^V]}{dt} = k [V^V][\text{aldehyde}][H^+]$$

at low acid concentrations, but the rate increased more rapidly at high acid concentrations than this equation would indicate. The rate of enolisation of these two aldehydes was found to be faster than the rate of oxidation and although the reaction kinetics are consistent with a reaction of the aldehyde in its enolic form, they cannot confirm this definitely. Various possible mechanisms for the reaction have been discussed.