[CONTRIBUTION FROM THE CHEMICAL LABORATORY, PRINCETON UNIVERSITY]

CATALYTIC OXIDATIONS IN AQUEOUS SOLUTIONS. II. THE OXIDATION OF PRIMARY ALCOHOLS

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Introduction

In a recent paper[2] of this series the author has shown that metallic oxides such as osmium tetroxide and vanadium pentoxide can be used, in aqueous solutions, to induce the oxidation of furfural and pyromucic acid by means of chlorates. The present paper describes several additional experiments in which vanadium pentoxide has been used to induce the oxidation of six of the primary alcohols.

The oxidation of alcohols has been the subject of a vast number of investigations which have occupied a period of more than a century. It is well established, therefore, that when primary alcohols are oxidized under ordinary conditions, the principal products which result from the oxidation are either aldehydes, or acids, or both. However, certain controlled oxidations of ethyl,[3] *n*-butyl,[4] *iso*butyl[5] and *iso*-amyl[6] alcohols yield esters as the principal products when the alcohols are present in excess. But such methods of oxidation are not applicable to all alcohols and any attempt to explain the formation of esters in these oxidations might lead to erroneous conclusions.

In the present paper a new catalytic method is described whereby the primary alcohols are simultaneously oxidized and esterified. It consists in the treatment of the alcohols, under proper care, with chlorates in dilute acid solution and in the presence of vanadium pentoxide as the catalyst.

Notwithstanding the great number of investigations published during the past hundred years on the oxidation of alcohols, comparatively little attention has been paid to the mechanism of the oxidation and subsequent esterification of these alcohols when the two reactions are taking place simultaneously in the same medium. Such a mechanism may be briefly postulated by the following equations

$$\mathrm{R{-}CH_2OH} \xrightarrow{\mathrm{O}} \mathrm{R{-}CHO + H_2O} \qquad (1)$$

$$\mathrm{R{-}CHO}\begin{array}{ll} +\ \mathrm{R{-}CH_2OH} \rightleftharpoons & \mathrm{R{-}CHOH(OCH_2R)} \\ & \qquad\quad \uparrow\!\downarrow \\ +\ 2\mathrm{R{-}CH_2OH} \rightleftharpoons & \mathrm{R{-}CH(OCH_2R)_2 + H_2O} \end{array} \qquad (2)$$

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[1] National Research Fellow in Chemistry.

[2] Milas, THIS JOURNAL, **49**, 2005 (1927).

[3] Godefroy, "Thèse de la Faculté des Sciences de Paris," **1886**, p. 43.

[4] Pierre and Puchot, *Ann.*, **163**, 283 (1872); *Ann. chim. phys.*, [4] **28**, 366 (1873).

[5] Robertson, "Organic Syntheses," John Wiley and Sons, Inc., New York **5**, 23 (1925).

[6] Balbiano, *Gazz. chim. ital.*, **6**, 238 (1876).

$$R—CHOH(OCH_2R) \xrightarrow{O} R—CO(OCH_2R) + H_2O \qquad (3)$$

It is quite obvious that the primary addition products of aldehydes when alcohols are present in excess would be unstable alcoholates or "hemiacetals," which may be regarded as secondary alcohols and therefore should be easily oxidized to the corresponding carbonyl derivatives before they are hydrolyzed. Although this assumption is purely hypothetical, at present, its justification is supported by the fact that, in the experiments to be described later, ester formation precedes acid formation. The organic acids formed during the reaction seem to result from the slow hydrolysis of the esters under the influence of the mineral acids. Furthermore, experiments in which acetaldehyde was used with the catalytic mixture failed to show an appreciable oxidation to acetic acid. However, when mixtures of equimolecular proportions of acetaldehyde and ethyl alcohol or acetaldehyde and *n*-butyl alcohol were used, instead of the aldehyde alone, ethyl and *n*-butyl acetates were formed, respectively.

In addition to hemiacetals, acetals might be formed as intermediate products of the oxidation of alcohols. Quite recently Adkins and his students[7] have carried out a series of experiments in which they found that the formation of acetals from mixtures of alcohols and aldehydes was markedly accelerated by certain salts and acids while the yields of acetals were greatly diminished with increased dilution. The latter observation of Adkins was more recently substantiated by Skrabal and Mirtl[8] who found that mineral acids act as catalysts in acetal hydrolysis which increases proportionately with increased dilution, while Skrabal, Brunner and Airoldi[9] assumed that such an hydrolysis takes place in steps hemiacetals being first formed. If this be the case, the acetals formed, during the oxidation of alcohols are immediately hydrolyzed under the influence of the mineral acids to hemiacetals which are then oxidized to the corresponding carbonyl derivatives. To confirm this view acetal, instead of ethyl alcohol, was used with the catalytic mixture. Ethyl acetate was the principal product of the oxidation.

Experimental Part

General Method.—The alcohols are oxidized very slowly, even at the boiling point of water, by a neutral aqueous mixture of sodium chlorate and vanadium pentoxide. However, in the presence of dilute sulfuric acid, the oxidation proceeds somewhat rapidly at first and slowly toward the end of the reaction. The appearance of a blue color, due to the reduction

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[7] (a) Adkins and Nissen, THIS JOURNAL, **44**, 2749 (1922); (b) Adams and Adkins, *ibid.*, **47**, 1358 (1925).

[8] Skrabal and Mirtl, *Z. physik. Chem.*, **111**, 108 (1924).

[9] Skrabal, Brunner and Airoldi, *ibid.*, **111**, 118 (1924).

of vanadium pentoxide to vanadium trioxide,[2] is a proof of the absence of chlorates and consequently of the end of the reaction.

After several preliminary experiments the following general method, which is applicable to alcohols higher than methyl alcohol, was adopted. Solutions of 5% sulfuric acid containing the calculated amount of sodium chlorate and 5–10% in excess necessary to oxidize half of the alcohol present completely to the corresponding acid were mixed with small amounts of vanadium pentoxide and the alcohol in question, in a suitable container provided with a reflux condenser and a dropping funnel, and the mixture was heated slowly to 75–80°. A rise of temperature above 80° at this stage of the reaction should be avoided since it causes a rapid evolution of chlorine dioxide, which forms an explosive mixture with the alcohols. During the progress of the reaction some of the sulfuric acid is used up and should be replaced by adding the calculated amount of 6 *N* sulfuric acid, assuming that all of the original acid is consumed after a heating of five hours. In some cases it was found necessary to make even a second addition of the acid. After the final addition, the temperature was raised to the boiling point of the mixture and kept there till the end of the reaction.

Method of Analysis.—The esters were hydrolyzed by refluxing each for two hours with an excess of about 0.5 *N* alcoholic potash solution. The excess alkali was titrated with a solution of 0.2 *N* hydrochloric acid, using phenolphthalein. Therefore, the results are expressed in terms of 0.2 *N* potassium hydroxide solution. A blank was run with each analysis to check the strength of the alkali.

The yields of organic acids were estimated by titration according to Richardson and Bowen.[10] For the sake of comparison, a blank was run with the estimation of each organic acid. The yields of both esters and organic acids, as expressed below, represent the results of several experiments performed with each alcohol.

I. Oxidation of Methyl Alcohol with Sodium Chlorate and Vanadium Pentoxide

The general method described above is not applicable to the oxidation of methyl alcohol because of the explosive violence of the reaction between the formaldehyde formed and the chlorine dioxide liberated during the reaction. Therefore, the following modification of the method was finally adopted.

Seventy-three g. of sodium chlorate (Merck's U. S. P. grade) was dissolved in 130 cc. of 2.5% sulfuric acid and the mixture was slowly dropped by means of a dropping funnel at the rate of about 1 cc. per minute into a flask fitted with an efficient reflux condenser and containing 64 g. of methyl alcohol (Merck's "c. p." grade, free from acetone), 50 cc. of water and 0.3 g. of vanadium pentoxide.[11] The mixture was heated to 75–80°. At first the solution becomes orange-yellow and changes to blue at the end of the reaction. After ten hours of heating, 36 cc. of 6 *N* sulfuric acid was slowly dropped into the mixture and heating continued until the blue color appeared. This took thirty-six hours longer. During the progress of the reaction a large quantity of carbon dioxide was evolved. At the end of the reaction the mixture was found to contain traces of

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[10] Richardson and Bowen, *J. Soc. Chem. Ind.*, **25**, 836 (1906).

[11] For the preparation of this catalyst, see Milas, ref. 2, p. 2007.

chloroform, as shown by the carbylamine test,[12] methyl formate, formic acid and small quantities of a condensation product the identity of which was not determined.

Isolation of Methyl Formate.—The reaction mixture from the above experiment was subjected to fractional distillation. The fraction boiling between 30 and 60° was collected, dried over anhydrous sodium carbonate and refractionated, the fraction distilling between 34 and 38° being collected and analyzed; yield, 5–6 g., or 8–10% of the alcohol taken.

Anal. Subs., 0.8372: required 68.8 cc. of 0.2 *N* KOH. Calcd. for $C_2H_4O_2$: 69.7 cc.

The aqueous solution was distilled once to free it from sulfuric acid and the blue color of vanadous oxide, which interferes with the end-point of the titration. The distillate was made up to a definite volume and lots of 50 cc. were titrated against 0.2 *N* sodium hydroxide (methyl orange, phenolphthalein). These indicated a yield of 8–9 g. of formic acid. No attempt was made to recover the unoxidized alcohol.

II. Oxidation of Ethyl Alcohol, Acetal and a Mixture of Equimolecular Proportions of Paraldehyde and Ethyl Alcohol with Sodium Chlorate and Vanadium Pentoxide

(*a*) **Oxidation of Ethyl Alcohol.**—Ninety-two g. of ethyl alcohol (abs.), 74 g. of sodium chlorate, 0.5 g. of the catalyst and 92 cc. of a 5% solution of sulfuric acid were brought together in a 500cc. Erlenmeyer flask provided with a reflux condenser and a dropping funnel. The mixture was heated very slowly to 75–80° and, after five hours of heating, 19 cc. of 6 *N* sulfuric acid was added very slowly and the temperature raised to the b. p. of the mixture. The blue color appeared after ten to twelve hours of heating. The mixture had separated into two layers, the upper of which was mostly ethyl acetate and some unconverted alcohol, while the lower contained alcohol, small quantities of acetaldehyde (forms a silver mirror in the cold with Tollens' reagent) and acetic acid. The entire mixture was subjected to fractional distillation; the fraction boiling between 70 and 90° was collected, dried over sodium carbonate, refractionated and the fraction distilling between 74 and 80° collected (74 g.) and analyzed by hydrolysis. This fraction proved to be 70–73% ethyl acetate, which is 60–61% of the total alcohol taken.

The entire aqueous solution, mixed with the fractions boiling above 80°, was distilled and the distillate made up to a definite volume. Samples of 50 cc. of this were titrated, indicating a total yield of 13.5–15.9 g. of acetic acid.

(*b*) **Oxidation of Acetal.**—One hundred g. of acetal (Kahlbaum's) was mixed with 62 g. of sodium chlorate, 0.3 g. of the catalyst and 100 cc. of 5% sulfuric acid and the mixture heated very slowly under a reflux condenser to 75–80°, this temperature being maintained till the end of the reaction, which took nine hours. The experiment yielded 45 g. of ethyl acetate, 36 g. of acetic acid and some alcohol which had presumably resulted from the hydrolysis of the ester.

(*c*) **Oxidation of an Equimolecular Mixture of Paraldehyde and Ethyl Alcohol.**—When acetaldehyde alone was treated with the oxidizing mixture in presence of 5% sulfuric acid, paraldehyde was chiefly formed, and when this mixture was further heated at 75–80° for thirty-four hours, no additional change was noticed. This mixture yielded only traces of acetic acid, as shown by the ferric chloride test.[13] However, when the aldehyde was mixed with equimolecular proportions of alcohol, oxidation occurred immediately, with ethyl acetate resulting as the main product of the reaction.

Fifty g. (1.14 moles) of paraldehyde (Kahlbaum's) was mixed with 52.3 g. (1.14 moles) of absolute alcohol, 50 g. of sodium chlorate, 0.3 g. of the catalyst and 100 cc. of

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[12] Hofmann, *Ann.*, **146**, 107 (1868).

[13] Weston, "A Scheme for the Detection of the More Common Classes of Organic Compounds," Longmans, Green and Co., **1912**, p. 27.

5% sulfuric acid and the mixture was heated very slowly, under a well-cooled reflux condenser, to 30°. In the beginning the temperature of the reaction mixture should not be allowed to rise above 40°, otherwise the reaction will become too violent and uncontrollable. After five hours the reaction has somewhat moderated and the temperature may now be raised to 75–80°, and 19 cc. of 6 *N* sulfuric acid slowly added to the mixture and the heating continued for six additional hours before the blue color appears. The experiment yielded 54 g. of ethyl acetate and 21.3 g. of acetic acid. Small quantities of unconverted alcohol and aldehyde were also recovered.

During the above reactions the active form of the aldehyde is presumably the monomolecular, which is regenerated under the influence of sulfuric acid. This is well shown by the work of McLeod[14] on the polymers of acetaldehyde.

These results, while not absolutely conclusive, indicate the plausibility of the view presented in the early part of this paper.

III. Oxidation of *n*-Propyl Alcohol with Sodium Chlorate and Vanadium Pentoxide

Ninety g. of *n*-propyl alcohol (Kahlbaum's) was mixed with 58 g. of sodium chlorate, 0.5 g. of the catalyst and 90 cc. of 5% sulfuric acid and the mixture slowly heated to 75–80° until the violent reaction moderated; then the temperature was raised to boiling. After five hours of heating, the first portion of 18 cc. of 6 *N* sulfuric acid was added, the heating continued for ten hours longer when a second portion of 22 cc. of the acid was slowly added, and the heating again continued until the blue color appeared. Twenty to twenty-two hours were required for the completion of the reaction. The mixture was then cooled and the ester layer separated from the aqueous layer. The former consisted chiefly of propyl propionate, propyl alcohol, propionic acid, small quantities of propionaldehyde and a chlorinated ester. It was shaken several times with a 4% solution of sodium hydroxide to remove the acid and most of the alcohol, washed twice with water, dried over anhydrous sodium carbonate and fractionated. The fraction distilling between 121 and 124° (42–45 g.) consisted of pure propyl propionate free from any chlorinated products. The ester yield was 48–51% of the total alcohol taken.

Anal. Subs., 1.2983: required 56.25 cc. of 0.2 *N* KOH. Calcd. for $C_6H_{12}O_2$: 55.96 cc.

A small quantity of the chlorinated ester boiling between 140 and 150° with decomposition was isolated but not analyzed.

The washings from the ester layer were mixed with the aqueous layer, treated with dilute sulfuric acid in excess and distilled. The distillate was made up to a liter, and a portion of it was exactly neutralized and evaporated to dryness on the water-bath. The dry salt, treated with absolute alcohol and a few drops of concd. sulfuric acid, gave a decided odor of ethyl propionate. Fifty-cc. samples of the distillate were then titrated and the total yield of propionic acid estimated. These indicated a total yield of 22–24 g. of propionic acid.

IV. Oxidation of *n*-Butyl Alcohol with Sodium Chlorate and Vanadium Pentoxide[15]

Eighty-nine g. of *n*-butyl alcohol (b. p. 116.5–118°) was mixed with 48 g. of sodium chlorate, 0.5 g. of the catalyst and 90 cc. of 5% sulfuric acid and the mixture heated

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[14] McLeod, *Am. Chem. Jour.*, **37**, 27 (1907).

[15] I am indebted to William L. Ruigh of this Laboratory for the preliminary experiments on this alcohol.

slowly, under a reflux condenser, to 75–80° for two hours, then to the b. p. for three hours longer, when 18 cc. of 6 *N* sulfuric acid was slowly added and the heating continued for five hours longer. A second portion of 22 cc. of the acid was then added and the heating continued till the blue color appeared. Fifteen to eighteen hours were required for the completion of the reaction. The mixture was then cooled and the layers separated. The ester layer was washed several times with a 4% solution of sodium hydroxide, twice with water, dried over sodium carbonate and fractionated. After two fractionations the fraction distilling between 164 and 166° (41–43 g.) was collected and analyzed; yield, 47–49% of the total alcohol taken.

Anal. Subs., 1.0892: required 38.22 cc. of 0.2 *N* KOH. Calcd. for $C_8H_{16}O_2$: 37.85 cc.

Besides the ester, small quantities of chlorinated products boiling with decomposition from 10–15° higher than the ester were isolated but not analyzed. From 20–25 g. of unoxidized alcohol was recovered. Small quantities of butylaldehyde were also isolated.

The washings from the ester layer were combined with the aqueous layer, treated with sulfuric acid and distilled, and the distillate was made up to a liter. A portion of this was exactly neutralized, and evaporated to dryness. The salt, treated with calcium chloride in the hot, formed a white precipitate which persisted only in cloudy form when the mixture was cooled. The odor of the free acid was characteristic of that of butyric acid. Fifty-cc. samples of the distillate were then titrated and indicated a yield of 5–6.5 g. of *n*-butyric acid.

V. Oxidation of *iso*Butyl Alcohol with Sodium Chlorate and Vanadium Pentoxide

Eighty-nine g. of *iso*butyl alcohol (Kahlbaum's) was mixed with 48 g. of sodium chlorate, 0.5 g. of the catalyst and 90 cc. of 5% sulfuric acid. The mixture was treated in exactly the same manner as in the case of normal butyl alcohol except that the number of hours to complete the reaction was only 15–16. The ester layer was also treated similarly, fractionated twice, and the fraction distilling between 147 and 150° (40–42 g.) collected and analyzed; yield, 46–48% of the alcohol taken.

Anal. Subs., 1.0282: required 36.15 cc. of 0.2 *N* KOH. Calcd. for $C_8H_{16}O_2$: 35.71 cc.

During the above oxidation large quantities of carbon dioxide were evolved, and some of the alcohol was oxidized to acetone, considerable quantities of which were detected by the iodoform reaction. *Iso*butylaldehyde and small quantities of chlorinated products were also isolated.

*Iso*butyric acid was identified by preparing its sodium salt, which was found to decolorize instantly dilute potassium permanganate. Acetone was identified as one of the oxidation products. Neither the salt nor the free acid formed a precipitate with calcium chloride either in the hot or in the cold. The total yield of *iso*butyric acid, as determined by titration, was 4–5 g.

VI. Oxidation of *iso*-Amyl Alcohol with Sodium Chlorate and Vanadium Pentoxide

Seventy-five g. of *iso*-amyl alcohol (Kahlbaum's pyridine free) was mixed with 40 g. of sodium chlorate, 0.5 g. of the catalyst and 150 cc. of a 5% solution of sulfuric acid and the mixture heated slowly to 75–80° for one hour; then the temperature was raised to boiling. After four hours 36 cc. of 6 *N* sulfuric acid was added and the heating continued for eight hours longer before the blue color appeared. The ester layer was shaken several times with a 4% solution of sodium hydroxide, washed twice with water, dried over sodium carbonate and fractionated. After three fractionations the fraction distilling between 190 and 195° (33–35 g.) was collected and analyzed; yield of *iso*-amyl *iso*-valerate, 45–48% of the alcohol taken.

Anal. Subs., 2.0247: required 59.40 cc. of 0.2 *N* KOH. Calcd. for $C_{10}H_{20}O_2$: 58.85 cc.

Besides *iso*-amyl *iso*valerate, *iso*valeric aldehyde, some unoxidized alcohol, small quantities of chlorinated products, acetone and carbon dioxide were identified.

When the washings from the ester layer were mixed with the aqueous layer and the mixture was treated with sulfuric acid, a layer of *iso*valeric acid separated out. The mixture was extracted with 300 cc. of ether and the ether extract dried over anhydrous sodium sulfate and subjected to fractionation. Most of the high-boiling liquid distilled between 175 and 177°. This was almost pure *iso*valeric acid; yield, 12–13 g. The acid had a very rancid odor and gave a gelatinous precipitate with a solution of zinc sulfate.

Summary

1. Vanadium pentoxide, in dilute sulfuric acid solution, induces the oxidation of methyl, ethyl, *n*-propyl, *n*-butyl, *iso*butyl and *iso*-amyl alcohols, acetal and mixtures of equimolecular proportions of paraldehyde and ethyl alcohol and paraldehyde and *n*-butyl alcohol with chlorates to give esters as the principal products.

2. Organic acids, small quantities of aldehydes and chlorinated products, acetone and carbon dioxide were isolated as by-products in the oxidation of some of the alcohols.

3. A tentative mechanism of the oxidation is presented.

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[Contribution from the Laboratory of Organic Chemistry of the University of Wisconsin]

HEMIACETAL FORMATION AND THE REFRACTIVE INDICES AND DENSITIES OF MIXTURES OF CERTAIN ALCOHOLS AND ALDEHYDES

By Homer Adkins and A. E. Broderick

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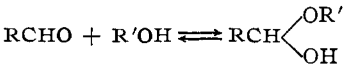
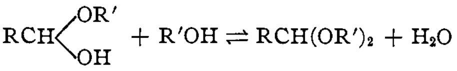
Heat is evolved when certain alcohols and aldehydes are mixed with each other. The addition of a catalyst such as hydrogen chloride or calcium chloride to the solution results in acetal formation. Presumably the pure alcohol and aldehyde react by addition according to the equation



$$RCHO + R'OH \rightleftharpoons RCH\begin{cases} OR' \\ OH \end{cases}$$

The hemiacetal in the presence of a catalyst may then react with alcohol to form an acetal and water



$$RCH\begin{cases} OR' \\ OH \end{cases} + R'OH \rightleftharpoons RCH(OR')_2 + H_2O$$