at the outlet of the reflux condenser. The solution was steam distilled until the distillate was neutral, and the total distillate was neutralized with excess barium hydroxide, filtered, and evaporated to dryness. The solid residue was treated with sulfuric acid, the barium sulfate was filtered, and 2 ml. of an ammonium sulfate solution was added to the filtrates, which were then subjected to paper chromatography with ethanol–ammonia solution–water (80:4:16) as solvent system. Comparisons were made with formic and acetic acids individually and together, in each case in the presence of ammonium sulfate. Both ninhydrin and the silver ion spray (for reducing substances) were employed. With ninhydrin spray, the R_f values for the formic and acetic acids both individually and as a mixture was 0.56 ± 0.02. The measured R_f value (ninhydrin) for the volatile acids from acetylalthiomycin was 0.56, the only other spot on the paper being due to ammonium sulfate near the origin. With basic silver ion spray, authentic acetic acid gave no spot, whereas formic acid or the acetic–formic acid mixture gave spots with an R_f value of 0.58. The volatile acids from acetylalthiomycin gave an R_f value of 0.58. These experiments demonstrate the presence of formic acid in the volatile acids from acetylalthiomycin.

Amino Acid Analyses of Althiomycin and Acetylalthiomycin.—The Moore and Stein method[19] for amino acid analysis was applied to hydrolysates of althiomycin and acetylalthiomycin.[20] These procedures, except for minor variations, were followed:

The sample (2 to 3 mg.) to be analyzed was hydrolyzed[19b] for 22 hr. in a sealed tube with redistilled 6 *N* hydrochloric acid (1 ml.). A constant temperature of 110° was maintained by suspending the tube above refluxing toluene. After cooling and drying in a vacuum desiccator over phosphorus pentoxide and sodium hydroxide, the sample was dissolved in water and 0.2 *M* sodium phosphate buffer, *p*H 6.5 (1 ml. total solution). The sample was then allowed to stand in air for 4 hours to oxidize cysteine to cystine before reacidification with *N* hydrochloric acid (0.06 ml.) and 0.2 *N* sodium citrate buffer, *p*H 2.2 (2 ml.). The resulting solutions were kept frozen until analyzed.

The most effective column material now in use is a strong cation exchange resin, Dowex 50-X4, hydraulically sieved, through 200 mesh. The pretreated resin[19a] as a slurry, resin–buffer (0.2 *N* sodium citrate–acetate, *p*H 5.09), was used to pack the water-jacketed column (0.9 × 150 cm.) under an air pressure of 10 cm. The column was washed overnight with 0.2 *N* sodium hydroxide under 10 cm. pressure. Before use, the column was equilibrated with 0.2 *N* sodium citrate buffer, *p*H 3.02, overnight until the *p*H of the eluate reached the *p*H of the developer. The column was mounted over a fraction collector actuated by an electronic drop counter calibrated to give 2-ml. fractions.

After introduction of the sample to the surface of the resin, the column was developed using the *p*H 3.02 buffer under 10-cm. pressure. The column was maintained at a temperature of 30° using a constant temperature bath with a water-circulating pump. The rate of elution was approximately 10 ml. per hour and was varied when necessary by a change in air pressure. After 21 to 24 hr. (210 to 240 ml.) the temperature was raised to 50°. A linear gradient system going from the 0.2 *N*, *p*H 3.02 buffer to a 2 *N*, *p*H 5.09 sodium citrate–acetate buffer (500 ml. of each) was started after 34 to 38 hours (340 to 380 ml.). A total of 1000 ml. of buffer was used to develop the columns.

The colorimetric detection procedure[19c] made use of the color reaction given by amino acids and ammonia with ninhydrin. The optical density of the solutions was determined using a Coleman Junior spectrophotometer at 570 mμ except for yellow solutions which were read at 440 mμ. The results given by various samples are presented in Table IV.

TABLE IV

RESULTS OF MOORE AND STEIN AMINO ACID ANALYSIS OF HYDROLYSATES

Sample[a]	Possible compound[b]	One residue recovd., %
Althiomycin (2.5 mg.)	Unknown (yellow)[c]	31
	Cystine[d,e]	37
	Ammonia[d]	217
	Unknown (red-violet)[f]	82
Althiomycin (2.47 mg.)	Cystine[d,e]	39
	Ammonia[d]	208
	Unknown (red-violet)[f]	95
Acetylalthiomycin (3.28 mg.)	Unknown (yellow)[c]	20
	Cystine[d,e]	26
	Ammonia[d]	115
	Unknown (red-violet)[f]	55
Oxidized acetylalthiomycin (2.23 mg.)	Cysteic acid[d]	70
	Unknown (yellow)[c]	6
	Ammonia[d]	236
	Unknown (red-violet)[f]	45

[a] Molecular weights assumed for purposes of calculation were: althiomycin, 722; acetylalthiomycin, 463; oxidized acetylalthiomycin, 511 (assumes that three oxygens were added). [b] Compounds listed in order of their elution. [c] Color yield presumed to be equal to that of proline, which produces a similar color. [d] Leucine was used as a standard reference, and correction was made for the varying color yield of each compound compared to leucine (ref. 19c). [e] Cysteine is air oxidized under the conditions of the experiment to cystine, which is analyzed as such. [f] Color yield based on leucine but not corrected to any amino acid relative to that of leucine.

All chromatograms showed many peaks which amounted to less than 10% of one residue. The unknown "yellow" peak gave the same color with ninhydrin as does proline, and appeared between the elution positions of proline and hydroxyproline. The unknown "red violet" peak was eluted where lysine would be expected. However, no evidence either by paper chromatography or electrophoresis could be obtained on these hydrolysates to support the possibility that lysine is present. The analytical procedure is particularly sensitive to contamination by ammonia, and the values obtained are undoubtedly high.

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(19) (a) S. Moore and W. H. Stein, *J. Biol. Chem.*, **211**, 893 (1954); (b) C. H. W. Hir, W. H. Steni and S. Moore, *ibid.*, **211**, 907 (1954); (c) S. Moore and W. H. Stein, *ibid.*, **211**, 941 (1954).

(20) The authors wish to thank Dr. J. G. Pierce for his generous help and for the use of his facilities in connection with these analyses.

[CONTRIBUTION FROM EMERYVILLE RESEARCH CENTER, SHELL DEVELOPMENT CO., EMERYVILLE, CALIF.]

The Oxidation and Reduction of Free Radicals by Metal Salts

BY HAROLD E. DE LA MARE, JAY K. KOCHI AND FREDERICK F. RUST

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The oxidation and reduction of organic free radicals by metal salts has been studied. The organic products of oxidation are highly dependent on the anion associated with the metal moiety. The use of such ligands as chloride, bromide and thiocyanate results in the formation of the corresponding substituted product, such as ethyl chloride from ethyl radical and cupric chloride. Sulfate or perchlorate salts yield products of elimination such as ethylene from ethyl radical and cupric sulfate. The mechanism of these processes is discussed; the former has been termed *ligand transfer* and the latter *electron transfer*. In both cases the metal moiety is reduced to the lower valence state. Oxy and thiyl radicals are unaffected by cupric salts which are so effective with carbon free radicals. Reduction of free radicals is possible by such metal ions as chromous, titanous and vanadous. In these reactions, carbanion intermediates are formed. Based on these reactions, preparative procedures for 5-hexenoic acid and esters, heptenoic acids, 6-substituted hexanoic acids, 6-heptene-2-one and methoxy-*t*-butoxybutenes are developed.

Introduction

The oxidation and reduction reactions of organic free radicals promoted by metal salts have been recognized by Waters for some time.[1] Recently, Dainton[2] and Bamford[3] and their co-workers have reported some quantitative values for the rates of some of these reac-

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(1) R. M. Haines and W. A. Waters, *J. Chem. Soc.*, 4256 (1955).

(2) E. Collinson, F. S. Dainton, D. R. Smith, G. J. Trudel and S. Tazuké, *Discussions Faraday Soc.*, **29**, 188 (1960).

(3) O. Bamford, A. Jenkins and R. Johnston, *Proc. Roy. Soc.* (London), **A239**, 214 (1957).

tions. Our investigations[4] have been primarily directed toward describing the types of reaction obtained between various free radicals and metal compounds. In this paper we wish to report the qualitative results of the oxidation and reduction reactions in solution of simple organic free radicals by metal salts.

Earlier papers have described the oxidation of free radicals by some metal salts, particularly ferric and cupric halides by a ligand transfer mechanism.[4b,d] In that process, reduction of the metal salt was postulated to occur by a direct transfer of the halide ligand to the free radical through a transition state such as I.

$$R\cdot + M^{+n}X_n \longrightarrow \underset{\text{I}}{[R\cdot\ X\text{-}MX_{n-1} \longleftrightarrow R\text{-}X\ MX_{n-1}]} \longrightarrow RX + M^{+n-1}X_{n-1}$$

This mechanism is analogous to the bridged activated complex described by Taube and co-workers[5] for the completely inorganic systems. In the absence of suitable bridging ligands, we have found that free radicals are oxidized under equivalent conditions by a competitive process which can be considered as involving electron transfer.

Results

Oxidation of Simple Radicals by Cupric Salts.—Alkyl radicals in solution were produced from several sources; these included reduction of *t*-alkyl hydroperoxides with ferrous salts, thermolysis of dialkyl peroxides and photolysis of ketones.

Methyl Radicals.—*t*-Butyl hydroperoxide reacted with ferrous sulfate or perchlorate in aqueous solution to produce *t*-butyl alcohol (45% yield) and acetone (53% yield) in addition to a gaseous mixture (52% yield based on peroxide) which consisted of methane (18%) and ethane (82%). These products arise from the reactions[6]

$$(CH_3)_3CO_2H + Fe^{+2} \longrightarrow (CH_3)_3CO\cdot + FeOH^{+2} \quad (1)$$

$$(CH_3)_3CO\cdot \longrightarrow CH_3COCH_3 + CH_3\cdot \quad (2)$$

$$CH_3\cdot + SH \longrightarrow CH_4 + S\cdot \quad (3)$$

$$2CH_3\cdot \longrightarrow C_2H_6 \quad (4)$$

SH = hydrogen donor

In the presence of excess ferric ion (3.0 *M*), 31% of the methyl radicals, which earlier appeared as methane and ethane, were oxidized to methanol and 43% appeared as methane and ethane. Cupric ion (0.1 *M*) under the same conditions oxidized 50% of the methyl radicals to methanol and only an 8% combined yield of methane and ethane was found. At higher (1.0 *M*) cupric ion concentrations, 66% of the methyl radicals were oxidized to methanol and less than 1% gaseous products were formed. If either cupric chloride or lithium chloride was added, the methyl radicals liberated in reaction 2 yielded methyl chloride in quantitative yields.

$$CH_3\cdot + Cu^{+2} \xrightarrow{H_2O} CH_3OH + Cu^+ \quad (5)$$

Since cuprous ions are readily oxidized by ferric ions, the reduction of hydroperoxides by a ferrous–cupric combination could conceivably be catalytic. However, the extraneous reduction of alkoxy radicals by ferrous ion,[7] in practice, disrupts the catalytic properties of the system. As expected, however, the ferrous–cupric pair is catalytically most efficient at low ferrous concentrations.

The results show that cupric ion is much more efficient as an oxidant than ferric ion. Moreover, cupric sulfate is approximately as effective as cupric or ferric chloride in oxidizing methyl radicals. The complexions of the products obtained from these two cupric salts in relatively dilute solutions differ entirely; with cupric chloride the product was exclusively methyl chloride, while in the absence of chloride ion, *i.e.*, sulfate anion, the product was largely methanol. The quantitative analysis for methanol in dilute aqueous solution by G.L.C. was difficult. The amount of methanol present relative to acetone indicated that 50% or more was not analyzable by G.L.C. It is possible that it was present as the water-soluble stable sulfate salt

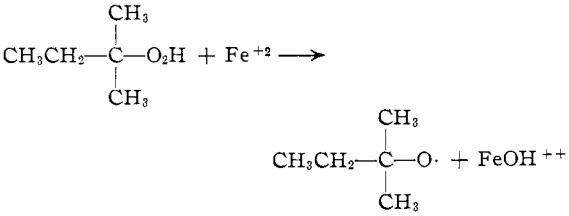
$$CH_3\cdot + CuSO_4 \longrightarrow CH_3OSO_3^- + Cu^+ \quad (6)$$

Functional group analysis of the crude reaction mixture showed the absence of other possible oxidation products, such as formaldehyde or formic acid. Carbon dioxide was not detected in the gas.

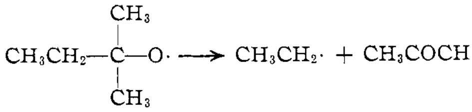
Cupric ion was also effective to a certain degree in the presence of such a relatively efficient hydrogen donor as isopropyl alcohol. Even in a 20% v. solution of isopropyl alcohol in water and 0.1 *M* cupric sulfate, less than 5% methane and ethane is formed. The rate of oxidation of methyl radicals by cupric ion is thus competitive with the chain transfer reaction.

Methyl radicals produced from the thermolysis of di-*t*-butyl peroxide in aqueous *t*-butyl alcohol solutions reacted with cupric ions in a similar manner. The photolysis of acetone in aqueous solution produced a gaseous mixture consisting of methane (85%), ethane (4%) and carbon monoxide (7%). In the presence of 0.5 *M* cupric sulfate, little methane was formed and methanol was the sole identifiable product. Quantitative results were obscured by the absorption of radiation by the cupric species.

Ethyl Radicals.—*t*-Amyloxy radicals from the reaction of *t*-amyl hydroperoxide and ferrous sulfate afforded ethyl radicals and acetone (85% yields based on peroxide).



$$CH_3CH_2C(CH_3)_2O_2H + Fe^{+2} \longrightarrow CH_3CH_2C(CH_3)_2O\cdot + FeOH^{++} \quad (7)$$



$$CH_3CH_2C(CH_3)_2O\cdot \longrightarrow CH_3CH_2\cdot + CH_3COCH_3 \quad (8)$$

Ethyl radicals reacted further to form a mixture (70%) of gaseous products which consisted of ethane (15%), ethylene (20%) and *n*-butane (65%), primarily by disproportionation and dimerization. In the presence of isopropyl alcohol, ethane was the predominant product (70–85%) with minor amounts of butane and ethylene. If cupric sulfate was added to the system initially, the gaseous product was ethylene (80% yield based on peroxide), contaminated with small amounts of butane and ethane. The latter contaminant can be eliminated if still higher cupric salt concentrations are employed. The most important liquid products were acetone (85%), *t*-amyl alcohol (10%) and ethyl alcohol (10%). The yields of the products which result from ethyl radical varied somewhat with cupric salt concentration.

When titanous chloride was employed as the reductant of the hydroperoxide the gaseous products were a mixture of ethane (80%) contaminated with small

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[4] (a) H. E. De La Mare, J. K. Kochi and F. F. Rust, *J Am. Chem. Soc.*, **83**, 2013 (1961); (b) J. Kumamoto, H. E. De La Mare and F. F. *Rust*, *ibid.*, **82**, 1935 (1960); (c) J. K. Kochi and F. F. Rust, *ibid.*, **84**, 3946 (1962) (d) J. K. Kochi, *ibid.*, **78**, 4815 (1956); **84**, 774, 1572, 1192, 2121, 2785 (1962); *Tetrahedron*, **18**, 483 (1962).

[5] H. Taube, *Advan. Inorg. Chem. Radiochem.*, **1**, 1 (1960).

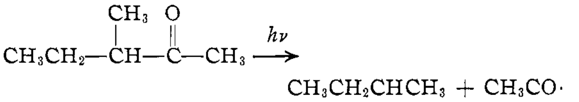
[6] A. Tobolsky and R. B. Mesrobian, "Organic Peroxides," Interscience Publishers, Inc., New York, N. Y., 1954, p. 95 ff.; R. J. Orr and H. L. Williams, *J. Am. Chem. Soc.*, **78**, 3273 (1956), and earlier papers.

[7] J. Kochi, *ibid.*, **84**, 1193 (1962).

amounts of *n*-butane, ethylene and ethyl chloride. When cupric sulfate was added to this system, the products from the ethyl fragment yielded ethyl chloride exclusively. Equivalent results were obtained from ferrous sulfate reductant if cupric chloride was added. In both cases only minor amounts of ethylene, ethane and *n*-butane were formed.

Diethyl ketone on photolysis[8] in aqueous solution yielded ethane (81%), ethylene (5%), *n*-butane (1%) and carbon monoxide (10%). The photolysis of a solution containing cupric sulfate produced a gaseous mixture containing ethylene predominantly. Similar results were obtained from the thermolysis of di-*t*-amyl peroxide in aqueous *t*-butyl alcohol solution.

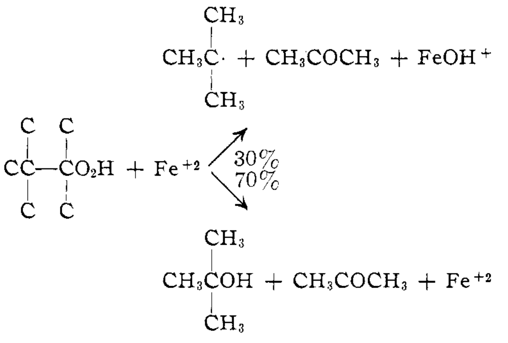
***sec*-Butyl and *t*-Butyl Radicals.**—The photolysis of 3-methyl-2-pentanone[8] in aqueous solution produced a gaseous mixture which consisted mainly of *n*-butane



$$\mathrm{CH_3CH_2{-}\overset{CH_3}{\overset{|}{C}}H{-}\overset{O}{\overset{\|}{C}}{-}CH_3 \xrightarrow{h\nu} CH_3CH_2CHCH_3 + CH_3CO\cdot} \qquad (9)$$

(56%), butylenes (27%) and ethylene (13%).[9] The butylene fractions contained butene-1 (66%), *trans*-butene-2 (27%) and *cis*-butene-2 (7%). A solution containing 0.5 M cupric sulfate under the same conditions yielded less butane (35%) and ethylene, but more butylenes (60%). The butylene fraction in this case consisted of butene-1 (32%), *trans*-butene-2 (44%) and *cis*-butene-2 (25%). The complexity of the photolysis process at present discourages the complete analysis of these results.[10] However, the increased yields of butylenes in the presence of cupric sulfate is in accord with expectations. The mole ratio of *sec*-butyl alcohol to butylenes which were formed from *sec*-butyl radicals by photolysis of the ketone in cupric solutions was 3.4.

Triptyl hydroperoxide (2,3,3-trimethyl-2-hydroperoxybutane) is extremely labile to acids, which induce a rearrangement to *t*-butyl alcohol and acetone.[11] In the presence of ferrous sulfate, triptyl hydroperoxide reacted by two competitive paths. The ionic rearrangements accounted for approximately 70% of the products, and



$$\mathrm{C\overset{C}{\underset{C}{C}}{-}\overset{C}{\underset{C}{C}}O_2H + Fe^{+2}} \begin{cases} \xrightarrow{30\%} \mathrm{CH_3\overset{CH_3}{\underset{CH_3}{C}}\cdot + CH_3COCH_3 + FeOH^+} & (10)\\ \xrightarrow{70\%} \mathrm{CH_3\overset{CH_3}{\underset{CH_3}{C}}OH + CH_3COCH_3 + Fe^{+2}} & (11)\end{cases}$$

ferrous reduction was responsible for the remaining 30%. By the latter path, a gaseous mixture was formed in 60% yield. It consisted of isobutylene (40%) and isobutane (60%). In the presence of 0.4 M cupric sulfate the liberated gas was exclusively isobutylene. Since the heterolytic reaction yielding *t*-butyl alcohol contributed significantly to the decomposition, it was not possible in this manner to assess the fates of the possible *t*-butyl carbonium ion intermediates.

The photolysis of pinacolone[8] in aqueous solution yielded a gaseous mixture consisting of isobutane (56%) and isobutylene (40%). In the presence of 0.15 M cupric sulfate the gas consisted mainly of isobutylene (85%) with a minor amount of isobutane (7%). The ratio of *t*-butyl alcohol: isobutylene formed in the latter reaction was 9.

Acetyl Radicals.—The photolysis of biacetyl produces mainly methane, acetaldehyde and carbon monoxide.[8] In the presence of cupric sulfate the photolysis yielded acetic acid, sulfuric acid and copper in quantitative yields according to the equation

$$\mathrm{CH_3COCOCH_3 + CuSO_4 \xrightarrow[H_2O]{h\nu} 2CH_3COOH + H_2SO_4 + Cu} \qquad (12)$$

When a cupric chloride solution of biacetyl in methanol was photolyzed, the cupric chloride was rapidly reduced to colorless cuprous chloride. The crude reaction mixture was directly distilled to remove the copper salts; distillate contained acetaldehyde, methyl acetate and methanol with a small amount of unidentified compounds. The yields of acetaldehyde and methyl acetate were approximately in the ratio of 2:1. Thus, in methanolic solutions, cupric chloride did not selectively oxidize acetyl radicals in competition with the much more readily available methyl alcohol solvent. In these reactions cupric chloride has only one-half the oxidative capacity of cupric sulfate, since cuprous chloride does not disproportionate to cupric salts and copper.

Acetyl radicals were also produced thermally from acetaldehyde and peroxide.[12] When acetaldehyde was treated in sealed tubes at 130° with di-*t*-butyl peroxide in aqueous *t*-butyl alcohol solutions containing cupric perchlorate, acetic acid was formed in 50-60% yields. Similar results were obtained in aqueous solution from the water-soluble sulfodi-*t*-butyl peroxide and sulfate. In the former reaction, *t*-butyl acetate was also isolated in approximately 10% yield.

Alkoxy Radicals.—The oxidations of several alkoxy radicals were examined. α-Cumyloxy radicals, from the reduction of cumene hydroperoxide by ferrous sulfate or thermolysis of dicumyl peroxide, yielded acetophenone, both in the presence and absence of cupric salts. Acetophenone is the product of the free radical decomposition of α-cumyloxy radicals.[13] If cumyloxy radicals were oxidized by cupric salts, among the expected products would be acetone and phenol.[14a,b]

Cyclohexyloxy radicals prepared from cyclohexyl hydroperoxide and ferrous sulfate yielded a mixture of cyclohexanol (23%) and cyclohexanone (77%). Cupric sulfate did not affect the relative yields of these products. Likewise, cyclohexenyloxy radicals reacted with *ferrous–cupric* sulfate to give a mixture of 2-cyclohexenone and 2-cyclohexenol (~77:23% w.).[15] The behavior of the cyclohexyloxy radicals would suggest that this reaction is independent of Cu^{++}; however, cyclohexenyl hydroperoxide was not treated with *ferrous ion* only. In a similar manner, *n*-butyl hydroperoxide reacted with ferrous sulfate to produce propane in 10% yield. In the presence of cupric sulfate, a 10% yield of propylene was formed under similar conditions. How-

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[8] E. W. Steacie, "Atomic and Free Radical Reactions," Reinhold Publishing Corp., New York, N. Y., 1954; Norrish, *et al.*, *J. Chem. Soc.*, 874 (1934); 1538 (1938).

[9] A. Nicholson, *Rev. Pure Appl. Chem.*, **2**, 174 (1952).

[10] The oxidation by cupric salts of butyl radicals from less ambiguous sources has been reported later. (J. K. Kochi, Abstracts 142nd National American Chemical Society Meeting, Atlantic City, N. J., 1962, p. 77-Q.)

[11] *Cf.* M. S. Kharasch, A. Fono and W. Nudenberg, *J. Org. Chem.*, **16**, 128 (1951).

[12] F. F. Rust and W. E. Vaughan, U. S. Patent 2,650,253 (to Shell Development Co.) (Aug. 25, 1953).

[13] M. S. Kharasch, A. Fono and W. Nudenberg, *J. Org. Chem.*, **16**, 113 (1951).

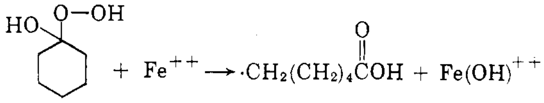
[14] (a) H. Hock and S. Lang, *Ber.*, **77B**, 257 (1944); (b) F. Seubold and W. E. Vaughan, *J. Am. Chem. Soc.*, **75**, 3790 (1953).

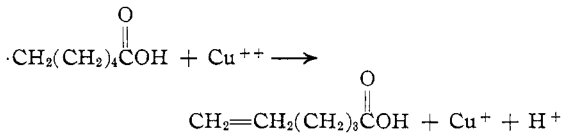
[15] Actually this ratio is only a rough approximation since it has not been corrected for small amounts of ketone and alcohol in the starting peroxide. Correction for the ketone in the peroxide indicated a yield of 2-cyclohexenone of ~58% (based on hydroperoxide).

ever, the relative yields of formaldehyde (10–15%) and butyraldehyde (45–55%) were unaltered by the presence of cupric salts.

Oxidation of Free Radicals as a Preparative Method

5-Hexenoic Acid.—The reaction of "cyclohexanone peroxide" with ferrous sulfate has been reported as a route to 1,12-dodecanedioic acid.[16a,b] In the reaction of the cyclohexanone peroxide with ferrous–cupric sulfate, this dibasic acid was essentially eliminated and *5-hexenoic* acid became the principal product (76% yield based on converted ketone and 61% based on hydrogen peroxide; functional group analyses). The formation of 5-hexenoic acid is rationalized as


(13)

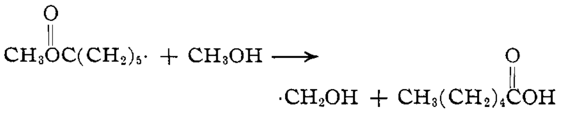

$$\cdot CH_2(CH_2)_4\overset{\overset{O}{\|}}{C}OH + Cu^{++} \longrightarrow CH_2{=}CH_2(CH_2)_3\overset{\overset{O}{\|}}{C}OH + Cu^{+} + H^{+} \quad (14)$$

$$Cu^{+} + Fe(OH)^{++} \longrightarrow Cu^{++} + Fe(OH)^{+} \quad (15)$$

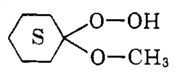
That reaction 14 does not proceed through a classical carbonium ion is evidenced by the relative unimportance of the 6-substituted hexanoic acids when the reaction was run in water or in methanol. The maximum formation of 6-hydroxyhexanoic acid from "cyclohexanone peroxide" in aqueous cupric–ferrous sulfate is estimated as ≤ 13% (based on converted ketone) under optimum conditions for producing the unsaturated acid.

Optimum yields of 5-hexenoic acid were not obtainable by working with catalytic quantities of metal salts. In fact, optimum yields were obtainable only by operating with relatively large amounts of ferrous sulfate (100 to 200% mole based on hydrogen peroxide) and somewhat smaller amounts of cupric sulfate (~50% mole). Ferrous ion concentration must be sufficiently high to allow reaction 13 to compete with reversion of the peroxide adduct to hydrogen peroxide, and cupric ion must be high enough to eliminate destruction of the radical by other paths. The ferrous ion consumption in a typical experiment (25 ± 5°) was of the order of 25 to 35 mmoles/100 mmoles of hydrogen peroxide reacted. As suggested previously, part of this loss is due to reaction with alkoxy radicals; the balance is probably due to oxidation by hydrogen peroxide.

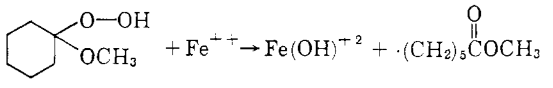
It is possible to prepare principally methyl 5-hexenoate by carrying out the same reaction in acidic aqueous methanol (43:57, v./v.). However, the ester and acid (58% and 8% yields, respectively, based on hydrogen peroxide) were comprised in part of saturated ester and acid (~31 mole % of the total product was saturated) produced by hydrogen abstraction from the solvent[17]


$$CH_3O\overset{\overset{O}{\|}}{C}(CH_2)_5\cdot + CH_3OH \longrightarrow \cdot CH_2OH + CH_3(CH_2)_4\overset{\overset{O}{\|}}{C}OH \quad (16)$$

Optimum yields of methyl ester required solubilizing cyclohexanone peroxide in acidic aqueous methanol prior to introduction of the metal salts. This was necessary to establish maximum concentration of the methyl hemiperketal

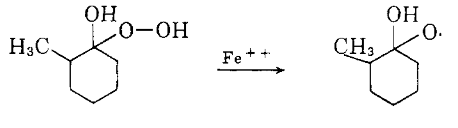

(17)

which is probably the intermediate reacting with ferrous ion to give the ω-methyl hexanoate radical


(18)

Although the synthesis of 5-hexenoic acid cannot be carried out with catalytic amounts of ferrous and cupric sulfate, it was possible to recycle the same metal salt solution through several batch reactions because of the low ferrous ion consumption per reaction pass.

Heptenoic Acids.—Methylcyclohexanone peroxide was treated with ferrous-cupric sulfate in the same manner described for cyclohexanone peroxide (*vide supra*). A mixture of heptenoic acids (~43% yield based on ketone) was produced and analyzed by a combination of ozonolysis, infrared and nuclear magnetic resonance spectroscopy. The production of a mixture of acids can be explained by the reactions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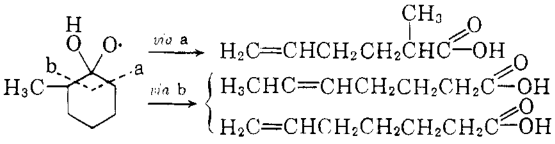
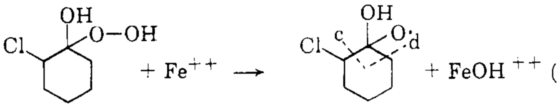

I II III (20)

Examination of a crude acid concentrate by nuclear magnetic resonance spectroscopy showed that compound I comprised ≤ 5% mole of the mixture. Quantitative infrared analysis and ozonolysis were in fair agreement with each other and indicated roughly 60% terminal and 40% internal double bond. Therefore, acids II and III were the principal products. Although the cleavage of the methylcyclohexyloxy radical appears to be almost cleanly on the methyl side, the conversion of the resultant radical to olefin is clearly not stereoselective.

Chlorohexenoic Acids.—A brief study of the reaction of the 2-chlorocyclohexanone peroxide with cupric–ferric sulfate was made. A poor yield (~11% based on ketone) of distilled acid was obtained; potentiometric titration indicated the distribution

	Equiv./100 g
"Stronger" carboxylic acid	0.14
"Weaker" carboxylic acid	0.52
Total	0.66
Theory for $C_6H_9O_2Cl$	0.74

The bulk of the chloroform-extractable product (~67% w.) appeared as a tarry kettle residue. Although evidence was obtained for the presence of carbonyl (VI) and acid (0.52 equiv./100 g.) in this residue, no positive identification was suggested. Chlorine analysis in-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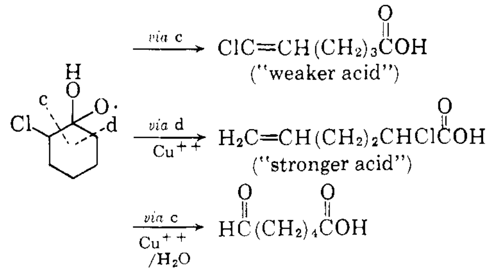
IV V VI (22)

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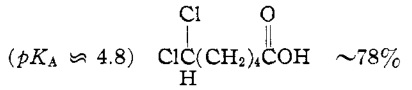
(16) (a) W. Cooper and W. H. T. Davison, *J. Chem. Soc.*, 1180 (1952); however, no yield data are reported here; (b) M. S. Kharasch and G. Sosnovsky, *J. Org. Chem.*, **23**, 1322 (1958).

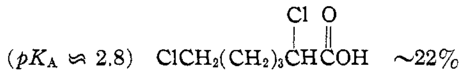
(17) In practice this synthesis suffers from the fact that the ester codistils with unreacted cyclohexanone upon simple distillation.

dicated that high-boiling chlorine-containing compounds were also present in this residue. Clearly, the quantitative aspects of the reaction are obscured by the low yield of distillable acid. Distribution of the distillable acids (*vide supra*) suggests the cleavage paths 21 and 22. Acids IV and V were formed in roughly a 4:1 ratio, respectively.

Additional support for the predominance of cleavage "c" came from treating the chlorocyclohexanone peroxide with ferrous sulfate and cupric chloride. The replacement of some sulfate with chloride was expected to give dichloro saturated acids and did.[4b] The mixture of chloroacids ($\sim$37% yield of distilled acid based on ketone) obtained by this technique was distributed as[18]



$$(pK_A \approx 4.8)\quad ClC\underset{H}{\overset{Cl}{|}}(CH_2)_4\overset{O}{\overset{\|}{C}}OH \quad \sim 78\% \qquad (23)$$

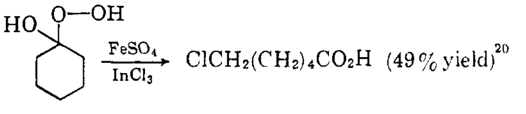


$$(pK_A \approx 2.8)\quad ClCH_2(CH_2)_3\overset{Cl}{\overset{|}{C}}H\overset{O}{\overset{\|}{C}}OH \quad \sim 22\% \qquad (24)$$

This distribution was obtained by nuclear magnetic resonance spectroscopy and was in good agreement with the potentiometric titration of the mixed acids. Thus, the conclusion is that cleavage at "c" is preferred over "d" by $\sim$4:1.

In general, synthetic procedures stemming from chlorocyclohexanone suffer by loss of the starting ketone to water-soluble by-products, the exact nature of which were not determined by these studies.

6-Substituted Hexanoic Acids.—Earlier ligand transfer work[4b] in our laboratory was extended to the ω-hexanoic acid radical. However, in the course of this study, we became aware for the first time of the independent work of Minisci and co-workers.[19] In summary, our work is in agreement with and corroborates their studies. Therefore, only limited mention will be made here of the preparation of ω-substituted hexanoic acids. Two syntheses of chlorohexanoic acid were conducted as



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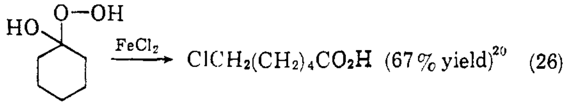
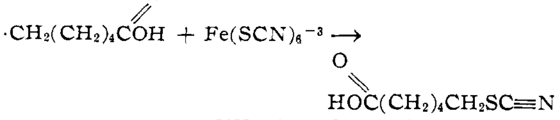


(26)

The ferric salt needed for the oxidation of the intermediate radical comes from the reduction of the cyclohexanone peroxide. Thus, as long as ferrous ion is available for reduction of the cyclohexanone peroxide, one can use any of a number of metallic chlorides; in fact, the inclusion of sodium or other alkali metal chloride is sufficient to give the chloro acid.

Another good bridging anion is thiocyanate, and the use of ferrous thiocyanate is illustrated by conversion of the ω-hexanoic radical to the 6-thiocyanatohexanoic acid



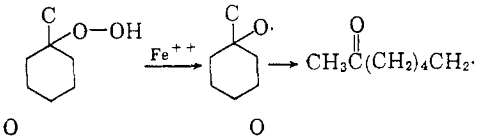
$$\cdot CH_2(CH_2)_4\overset{O}{\overset{/\!/}{C}}OH + Fe(SCN)_6^{-3} \longrightarrow HO\overset{O}{\overset{\backslash\!\backslash}{C}}(CH_2)_4CH_2SC{\equiv}N \qquad (27)$$

VII ($\sim$64% yield)[20]

Isolation of VII was carried out as the methyl ester. The ester was isomerized presumably to the isothiocyanate by heating; however, rigorous identification of the latter was not made.

The ω-hexanoic acid radical is a very convenient source of 6-substituted derivatives, and theoretically whenever the reaction is run in the presence of a "bridging anion" (X^-, SCN^-, N_3^-, CN^-, $S_2O_3^=$, $^-S\overset{\overset{S}{\|}}{C}OR$, etc.), that specific anion may be attached to the ω-position in reasonably good yield. The extensive work of Minisci and co-workers illustrates many synthetic possibilities in this field.[21]

6-Hepten-2-one.—The reaction of alkyl-substituted cyclic hydroperoxides with metal ions[22] offers a convenient route to bifunctional compounds. For example, methylcyclohexyl hydroperoxide reacted with ferrous–cupric sulfate to produce excellent yields of 6-hepten-2-one (65 to 94%). The synthesis of this ketoölefin occurs *via* the reaction scheme



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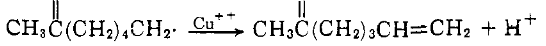
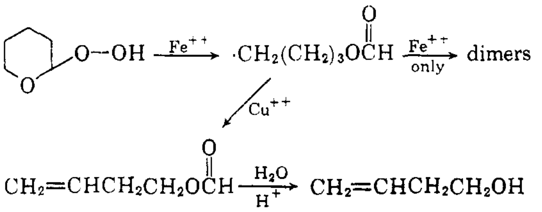


$$CH_3\overset{O}{\overset{\|}{C}}(CH_2)_4CH_2\cdot \xrightarrow{Cu^{++}} CH_3\overset{O}{\overset{\|}{C}}(CH_2)_3CH{=}CH_2 + H^+ \qquad (29)$$

Clearly, where a given cyclic hydroperoxide is available or easily synthesized, the synthesis of the unsaturated carbonyl becomes an easy one-step process therefrom.

1-Butenol.—The reaction of 2-tetrahydropyranyl hydroperoxide with ferrous sulfate has been shown to give good yields of 1,8-octanediformate.[23] However, in the presence of cupric–ferrous sulfate the cyclic peroxide was converted in modest yield ($\sim$20%) to 3-butene-1-ol. The latter product is believed to result from the hydrolysis of the intermediate 3-butenyl formate



(30)

Although the over-all product distribution could be changed by the manner of mixing the peroxide and metal ions (see Experimental), it was not possible in our hands to produce better than $\sim$20% yields of 3-butene-1-ol. The balance of the peroxide appeared largely as high boiling formate ester plus some non-formate ester (presumably δ-valerolacetone). The fact that the ω-butyl formate radical cannot be intercepted by cupric ion to give the 3-butenyl formate (or "ol") in a yield comparable to that obtained in dimerization by ferrous ions ($\sim$56%) is puzzling and suggests that the parent alkoxy radical may be diverted by competitive metal ion reactions. The exact nature of this diversion has not been elucidated, but previous work cited suggests that direct oxidation of this alkoxy radical by cupric ion is not likely.

Methyl Ethers of *t*-Butoxybutenols.—Previous workers[24] have demonstrated the addition of *t*-butoxy

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[18] We are indebted to Dr. C. A. Reilly of this Laboratory for the n.m.r. analysis of these chloro acids.

[19] F. Minisci, *Angew. Chem.*, **70**, 599 (1958).

[20] Based on hydrogen peroxide.

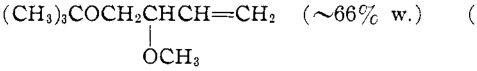
[21] F. Minisci and U. Pallini, *Gazz. chim. ital.*, **89**, 2438 (1959); F. Minisci, *ibid.*, **89**, 2428 (1959); F. Minisci and A. Portolani, *ibid.*, **89**, 1922, 1941 (1959); F. Minisci, *ibid.*, **89**, 626, 1910 (1959).

[22] See E. G. E. Hawkins and D. P. Young, *J. Chem. Soc.*, 2804 (1950), for discussion of reaction with ferrous sulfate only.

[23] Unpublished work by G. M. Coppinger, H. E. De La Mare, E. A. Youngman and F. F. Rust of this Laboratory. See, however, N. Milas, R. Peeler and O. Mageli, *J. Am. Chem. Soc.*, **76**, 2323 (1954).

[24] M. S. Kharasch, F. S. Arimoto and W. Nudenberg, *J. Org. Chem.*, **16**, 1556 (1954).

radicals to butadiene and subsequent dimerization of the intermediate radical in the presence of ferrous sulfate only. When the reaction of butadiene (excess) with *t*-butyl hydroperoxide was carried out with ferrous–cupric sulfate in methanol–water (30:26 v./v., 0 to −3°), good yields (55–60%) of the methyl ethers of isomeric *t*-butoxybutenols were produced with the distribution



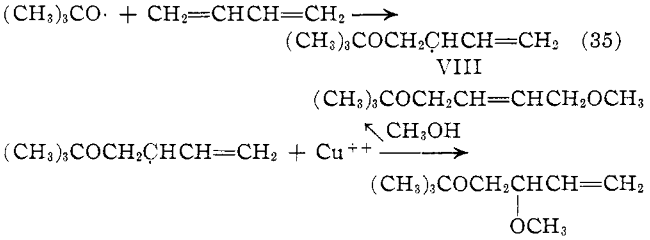
$$(CH_3)_3COCH_2\underset{\displaystyle OCH_3}{\underset{|}{C}}HCH{=}CH_2 \quad (\sim 66\%\ w.) \qquad (31)$$

$$trans^{25}\text{-}(CH_3)_3COCH_2CH{=}CHCH_2OCH_3\ (\sim 32\%\ w.) \qquad (32)$$

$$(CH_3)_3COCH_2CHOHCH{=}CH_2\ \ (\sim 2\%\ w.)(?) \qquad (33)$$

The products are accounted for on the basis of the sequence

$$(CH_3)_3COOH \xrightarrow{Fe^{++}} (CH_3)_3CO\cdot \qquad (34)$$

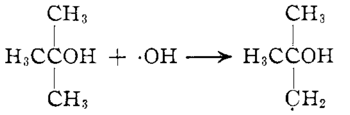


$$(CH_3)_3CO\cdot + CH_2{=}CHCH{=}CH_2 \longrightarrow \underset{\text{VIII}}{(CH_3)_3COCH_2\dot{C}HCH{=}CH_2} \qquad (35)$$

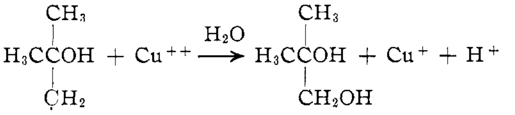
$$(CH_3)_3COCH_2\dot{C}HCH{=}CH_2 + Cu^{++} \xrightarrow{CH_3OH} (CH_3)_3COCH_2CH{=}CHCH_2OCH_3 \ \text{and}\ (CH_3)_3COCH_2\underset{\displaystyle OCH_3}{\underset{|}{C}}HCH{=}CH_2$$

Cursory experiments were also carried out in a urea–water mixture,[25,26] and as expected in the presence of cupric ion dimerization of VIII was prevented. However, in the presence of urea the bulk of the product appeared to be a urea adduct formed in the manner shown for methanol. In a similar manner, the ω-hexanoic acid radical was also added to butadiene to give the methoxy-substituted decenoic acid and/or methyl esters.[27]

Isobutylene Glycol.—The reaction of hydroxyl radicals ($H_2O_2 + Fe^{++}$) with *t*-butyl alcohol to give 2,5-dimethyl-2,5-hexanediol has been well established.[28] When this reaction was carried out in the presence of cupric–ferrous sulfate, evidence was obtained for the production of α-glycol (0.36 mole/mole H_2O_2)



$$H_3C\overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle CH_3}{\underset{|}{C}}}}OH + \cdot OH \longrightarrow H_3C\overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle \dot{C}H_2}{\underset{|}{C}}}}OH \qquad (37)$$



$$H_3C\overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle \dot{C}H_2}{\underset{|}{C}}}}OH + Cu^{++} \xrightarrow{H_2O} H_3C\overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle CH_2OH}{\underset{|}{C}}}}OH + Cu^{+} + H^{+} \qquad (38)$$

A single attempt to isolate the glycol from the acidic system resulted in the production of isobutyraldehyde and its condensation products. Nevertheless, this demonstrates another synthetic possibility for radical trapping with cupric ion. In general, the hydroxyl radical–cupric ion couple would appear to be capable of reacting with other substrates in which hydroxyl radical gives good yields of radical dimer.

Reduction of Free Radicals by Metal Salts.—In addition to the transfer of an electron *from* the free radical to the metal salt, it is possible to effect transfer of an electron *to* the free radical from the metal salt to generate a carbanion. The oxidation of metal

$$CH_3\cdot + M^{+n} \longrightarrow CH_3\colon + M^{+n+1} \qquad (39)$$

$$CH_3\colon + D_2O \longrightarrow CH_3\text{–}D + D^- \qquad (40)$$

salts by free radicals was qualitatively investigated with methyl radicals in deuterium oxide. Methyl carbanions produced from the reduction of the free radical under these conditions was identifiable as deuteriomethane mass spectroscopically. In blank experiments methyl radicals were generated in the absence of metal ions by thermolyzing di-*t*-butyl peroxide in isopropyl alcohol–water solutions containing deuterium oxide. The absence of deuteriomethane in the gaseous product (methane) indicated that the methyl radicals did not react with water or its exchange products (acetone and isopropyl alcohol). In the same solvents, methyl radicals were produced from *t*-butyl hydroperoxide and reducing agents listed in Table I. From the relative amounts of methane and deuteriomethane one can calculate the number of methyl radicals which react by hydrogen chain transfer with isopropyl alcohol and those which react by metal salt reduction. These preliminary results do not have quantitative significance other than to denote the relative reducing abilities of certain metal salts.

Table I

Reduction of Methyl Radicals by Metal Salts[a]

Metal salt	CH_4/CH_3D
$FeSO_4$[b]	2100
$TiCl_3$[b]	32
$TiCl_3$[c]	53
$TiCl_3$[d]	18
V_2SO_4	11
$CrSO_4$	4.6

[a] In solution consisting of 50% v. isopropyl alcohol, 10% v. D_2O, 40% H_2O. [b] *t*-BuOH added to metal salt solution. [c] Metal salt solution added to hydroperoxide. [d] No isopropyl alcohol.

Discussion

The use of metal salts as oxidants of free radicals is highly dependent on the nature of the anion associated with the metal moiety. For example, cupric chloride and bromide oxidize alkyl radicals to the corresponding alkyl halides, but cupric sulfate under similar conditions yields unsaturated products as well as substitution products. In competition experiments it was determined that the rate of the ligand transfer reaction is faster than that of the latter. Both reactions, however, are qualitatively much faster than the usual reaction of free radicals such as chain transfer, addition, etc. Thus, in practice, these salts inhibit free radical chain processes. Either cupric sulfate or perchlorate in aqueous solution oxidizes such primary radicals as ethyl and 6-hexanoic acid radicals to ethylene and 5-hexenoic acid, respectively, in good yields. Minor amounts of the substitution products ethanol and 6-hydroxyhexanoic acid are also formed. In alcohol solvents, ethers result in place of alcohols. With cupric chloride under conditions similar to those above, ethyl and 6-hexanoic acid radicals are converted cleanly to ethyl chloride and 6-chlorohexanoic acids, respectively. The elimination products in these cases are only minor contaminants.

As prototypes cupric chloride and cupric sulfate represent two characteristic oxidants of free radicals by metal salts. The oxidation of free radicals by metal salts such as cupric chloride to yield substitution products has been previously described as *ligand transfer process*.[4b,d] We wish here to represent the oxidation of free radicals by other metal salts such as cupric sulfate (which yield elimination products) as an *electron transfer process*.

Among various free radical types, the electron transfer process is most significant with primary alkyl radicals. Thus, the high yields of ethylene from ethyl radicals are not paralleled by the formation of isobutylene from *t*-butyl radicals. In general, the partitioning of free radicals between elimination and substi-

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(25) The presence of some *cis* cannot be precluded.

(26) The reaction was run using urea as an antifreeze; see ref. 24.

(27) J. Kochi and F. F. Rust, *J. Am. Chem. Soc.*, **84**, 3946 (1962).

(28) D. D. Coffman, E. L. Jenner and R. D. Lipscomb, *ibid.*, **80**, 2865 (1958); see also U. S. Patent 2,700,051.

tution products correlates roughly with the ionization potentials of the free radicals obtained in the gas phase[29] as shown in Table II. Thus, the amount of substitution product relative to elimination product from the oxidation of free radicals by cupric sulfate increases roughly as the ionization potential of the radical decreases. Or as a corollary, it can be stated that the *more stable the radicals or the incipient carbonium ion formed, the more probable that the products will be those expected from the solvolysis of a classical carbonium ion.* Thus, while primary alkyl radicals are oxidized by cupric sulfate in aqueous solution mainly to olefins, tertiary radicals are oxidized predominantly to alcohols, and allylic radicals are oxidized exclusively to alcohols. The same oxidation of tertiary or allylic radicals in methanol affords the corresponding methyl ethers. The formations of ethers and alcohols from these protic solvents appear to result from common or related intermediates.

TABLE II

OXIDATION OF FREE RADICALS BY CUPRIC SULFATE

Radical	Cupric ion concn., *M*	Olefin yield, %	Alcohol yield, %	ΔH_f,[a] kcal./mole
Methyl[c]	0.05	0	81	262
Ethyl[b]	.15	68	19	224
sec-Butyl[c]	.05	19	65	180
t-Butyl[c]	.05	9	82	166
Allylic[d]	.05	0	85	218
Acetyl[c]	.05	0	100	163–203

[a] Heat of formation of the gaseous ion.[29] [b] From *t*-alkyl hydroperoxide and Ti^{+3}. [c] From photolysis of the methyl ketone. [d] From *t*-butoxy radical and butadiene.

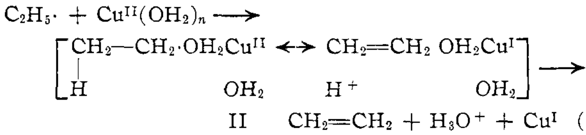
Since the heats of hydration of these gaseous carbonium ions have not yet been determined, it is not possible to pursue this correlation too vigorously. Thus, the ionization potentials given in Table II are not in complete accord with the correlation stated previously. Allylic radicals, for example, give more substitution products than either *sec*-butyl or *t*-butyl radicals, although gaseous ionization potential is higher than that of either of the latter radicals. Qualitatively, however, it is expected that the resonance-stabilized allylic carbonium ions would exhibit different solvation requirements from that of its saturated counterparts, and the deviations may be attributed to the differences in the energies of solvation of the carbonium ions. It has been recognized from studies of solvolytic reactions[30] that allylic carbonium ions are incipiently more readily formed than either secondary or tertiary alkyl carbonium ions, and the differences in rates have been attributed to the greater thermodynamic stability of the allylic systems. Entropy considerations have not been made.

Formally there are several mechanisms which can be postulated for the oxidation of free radicals by cupric salts. Although the difficult kinetic measurements have not yet been made, the contrary notwithstanding it will be assumed that the oxidation is a bimolecular reaction of a free radical and a cupric salt. In the two-step process the oxidation is postulated to proceed *via* a carbonium ion intermediate which results from the transfer of the electron from the radical to the cupric salt. This intermediate is similar in type to those postulated in the nitrous acid deamination of amines[31] and the solvolytic reactions of halides and esters.[30] The relative yields of elimination products from the later reactions are lower than those obtained from the oxidation of free radicals by metal salts. For example, ethylamine on treatment with nitrous acid in aqueous solutions affords ethylene and ethanol in 5 and 38% yields,[32] respectively, whereas the oxidation of ethyl radicals by cupric sulfate under similar conditions yields ethylene and ethanol in 68 and 19% yields, respectively. It is possible that carbonium ions formed from the nitrous acid deamination reactions are not completely free of the nitrogen[31b] or are solvated to a different degree than those formed from the oxidation of free radicals. Similarly, the solvolysis of primary halides and esters produces only small amounts of olefins.[30] In either case it appears that carbonium ions formed by an electron transfer process show a unique predisposition toward elimination reactions relative to those resulting from conventional carbonium ion processes.

$$CH_3CH_2\cdot + Cu^{II} \longrightarrow CH_3CH_2^+ + Cu^I \qquad (41)$$

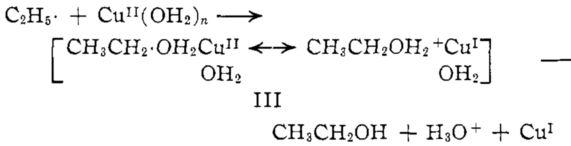
$$\begin{tikzcd} & CH_2{=}CH_2 + H_3O^+ \\ CH_3CH_2^+ \arrow[ur] \arrow[dr] \arrow[r, phantom, "H_2O"] & \\ & CH_3CH_2OH_2^+ \end{tikzcd} \qquad (42)$$

An alternative mechanism to that which postulates the formation of a carbonium ion intermediate is one in which the reaction between the free radical and cupric salt occurs by two concurrent reactions proceeding through *similar* transition states. In aqueous solutions the transition states for these one-step processes can be formulated as II and III



$$C_2H_5\cdot + Cu^{II}(OH_2)_n \longrightarrow \left[\begin{matrix} CH_2{-}CH_2\cdot OH_2Cu^{II} \\ | \qquad\qquad\quad OH_2 \\ H \end{matrix} \longleftrightarrow \begin{matrix} CH_2{=}CH_2\ OH_2Cu^{I} \\ H^+ \qquad\qquad OH_2 \end{matrix}\right] \longrightarrow CH_2{=}CH_2 + H_3O^+ + Cu^I \qquad (43)$$

II



$$C_2H_5\cdot + Cu^{II}(OH_2)_n \longrightarrow \left[\begin{matrix} CH_3CH_2\cdot OH_2Cu^{II} \\ OH_2 \end{matrix} \longleftrightarrow \begin{matrix} CH_3CH_2OH_2{}^+Cu^{I} \\ OH_2 \end{matrix}\right] \longrightarrow CH_3CH_2OH + H_3O^+ + Cu^I \qquad (44)$$

III

Transition state II differs from the electron transfer reaction described earlier in that the β-hydrogen is lost simultaneously with electron transfer and no carbonium ion intermediate is formed. In protic solutions such as water and alcohol the place and number of solvent molecules in the locus of the electron transfer milieu is as yet difficult to assess and not only the number of solvent molecules but also the molecular displacements of the reacting species may differ in the transition states II and III. Transition state III, in which the solvent constitutes the bridging group, is formally analogous to ligand transfer processes generally.

In both mechanisms, related transition states are involved in the elimination and substitution processes. The subtle variation in the molecular geometry of the transition states leading to the several products is reminiscent of the merged bimolecular substitution and elimination reactions described by Winstein and coworkers for solvolytic reactions.[33]

The choice between the discrete carbonium ion mechanism and the synchronous one-step processes is at present a difficult and a moot one. In essence these paths differ by the extent to which the β-hydrogen bond is broken in the rate-limiting step. In the former case this bond is largely intact throughout the oxidation step, while in the latter a modicum of the breaking of the β-hydrogen bond accompanies the electron transfer from the radical to the metal ion.

There is evidence that both of these mechanisms are operative to a certain extent in individual free radical

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[29] F. Field and J. Franklin, "Electron Impact Phenomena," Academic Press, Inc., New York, N. Y., 1957.

[30] A. Streitwieser, Jr., *Chem. Rev.*, **56**, 571 (1956).

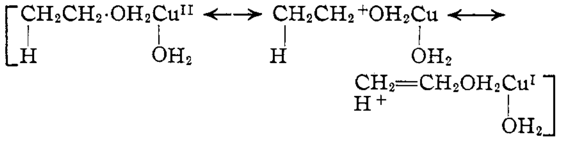
[31] (a) DeWolfe and W. G. Young, *ibid.*, **56**, 753 (1956). (b) A. Streitwieser, Jr., *J. Org. Chem.*, **22**, 861 (1957).

[32] (a) J. D. Roberts and J. A. Yancey, *J. Am. Chem. Soc.*, **74**, 5943 (1952); (b) F. C. Whitmore and R. S. Thorpe, *ibid.*, **63**, 1118 (1941).

[33] S. Winstein, D. Darwish and N. Holness, *ibid.*, **78**, 2915 (1956).

cases. The carbonium ion mechanism is most satisfactory with certain free radicals which produce relatively stable carbonium ions. The oxidation of tertiary alkyl radicals, benzylic radicals and allylic radicals to yield mainly substitution products is in harmony with the postulation of carbonium ion intermediates. Similarly there is no evidence for ketene intermediates in the oxidation of acyl radicals. The *t*-butyl radicals formed from several sources show a pronounced propensity for forming primarily products which correspond to the solvolytic reaction of the *t*-butylcarbonium ion, together with some elimination product, isobutylene. This rationale, however, is difficult to apply to primary alkyl systems like ethyl radical. The high yield of elimination product is not consistent with the postulation of an intermediate carbonium ion, at least not of the type usually encountered in organic reactions. Although some of the solvolytic product, ethanol, is formed, the appearance of ethylene as a major product is contemporarily taken as evidence against the formation of a free ethyl carbonium ion in aqueous solution.[34] Good yields of ω-hexenoic acid are also obtained from the reaction of the cyclohexanone hydrogen peroxide adduct with ferrous sulfate in the presence of cupric sulfate. The oxidation of primary radicals to afford unsaturated products in high yields has no analogy in classical carbonium ion reactions.

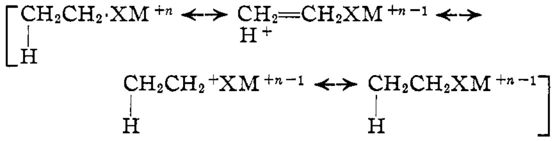
Whether the oxidation of free radicals by cupric salts proceeds by a one-step concerted or a two-step carbonium ion mechanism, it is clear from the reactivities of free radicals that a sizable amount of positive charge is incorporated by the organic moiety in the transition states. Thus, oxy and thiyl radicals are unaffected by cupric salts. Even such radicals as *n*-butoxy and cyclohexyloxy which possess β-hydrogens do not appear to be oxidized by cupric salts. The inability of cupric salts to oxidize even alkoxy radicals which possess removable β-hydrogens indicates that a considerable driving force for the oxidation is derived from the removal of the electron from the free radical center. The ionization potential of oxygen (13.64 v.)[35] is greater than that of carbon (11.26 v.)[35] and the transfer of an electron from an oxy radical would be expected to be more difficult than that from an analogous carbon center. Moreover, the selectivity is not restricted to oxy and thiyl radicals since other experiments[27] have shown that negatively substituted carbon radicals show different reactivities toward various cupric salts. Thus the determining factor in an oxidation of a free radical is the ease with which the system can achieve cationic character in the transition state, and the contribution to the driving force from the loss of a β-proton cannot be very large.[36] Since there appears to be a significant component of electron transfer even in the one-step mechanism for oxidative elimination, the transition state II is more properly represented as II′.



$$\left[\underset{\text{H}}{\text{CH}_2}\text{CH}_2\cdot\text{OH}_2\underset{\text{OH}_2}{\text{Cu}^{\text{II}}} \longleftrightarrow \underset{\text{H}}{\text{CH}_2}\text{CH}_2{}^{+}\text{OH}_2\underset{\text{OH}_2}{\text{Cu}} \longleftrightarrow \underset{\text{H}^+}{\text{CH}_2}\text{=}\text{CH}_2\text{OH}_2\underset{\text{OH}_2}{\text{Cu}^{\text{I}}}\right] \quad (45)$$

II′

Finally the dichotomy of these oxidative processes of free radicals into discrete categories is arbitrary. The distinction of the mechanism as a two-step carbonium ion, a one-step oxidative elimination or a one-step ligand transfer reaction may be too arbitrary. It is conceivable that all of these reactions proceed through quite related transition states in which the resonance contributions from electron transfer and ligand transfer as indicated in IV



$$\left[\underset{\text{H}}{\text{CH}_2}\text{CH}_2\cdot\text{XM}^{+n} \longleftrightarrow \underset{\text{H}^+}{\text{CH}_2}\text{=}\text{CH}_2\text{XM}^{+n-1} \longleftrightarrow \underset{\text{H}}{\text{CH}_2}\text{CH}_2{}^{+}\text{XM}^{+n-1} \longleftrightarrow \underset{\text{H}}{\text{CH}_2}\text{CH}_2\text{XM}^{+n-1}\right] \quad (46)$$

IV

vary not only with the structure of the free radical, but also with the identity of the metal species and counterions. Further experiments are in progress which will stress these points.

Experimental

Materials.—Metal salts: **ferrous sulfate** ($FeSO_4 \cdot 7H_2O$), Baker and Adamson reagent; **cupric sulfate** ($CuSO_4 \cdot 5H_2O$), Baker and Adamson reagent; **cupric chloride** ($CuCl_2 \cdot 2H_2O$), Baker and Adamson reagent; **ferric ammonium sulfate** [$Fe_2(NH_4)_2(SO_4)_4 . 24H_2O$], Baker and Adamson reagent; **titanous chloride** ($TiCl_3$ aqueous solution, 20%), Fisher Scientific Co.; **cupric perchlorate**, aqueous solution prepared by neutralizing **cupric carbonate** (Baker analyzed, J. T. Baker Co.) with **perchloric acid** (71%, Baker and Adamson reagent) potentiometrically; **indium chloride**, Indium Corp. of America (aqueous solution).

Peroxides: hydrogen peroxide, 30% analytical reagent, Mallinckrodt Chemical Co.; 90% commercial grade, Shell Chemical Co.; **di-*t*-butyl peroxide**, Shell Chemical Co.; ***t*-butyl hydroperoxide**, Lucidol Corp., 65% purified by vacuum distillation, m.p. 5°; **cumene hydroperoxide**, Hercules Powder Co., purified through sodium salt[14b]; ***t*-amyl hydroperoxide**, from *t*-amyl alcohol and hydrogen peroxide[37]; ***n*-butyl hydroperoxide**, from *n*-butyl methanesulfonate and hydrogen peroxide[38]; **methyl cyclohexyl hydroperoxide**, synthesized from methyl cyclohexyl alcohol[39] which, in turn, was synthesized by a conventional Grignard reaction; **2-tetrahydropyranyl hydroperoxide**, synthesized from dihydropyran and hydrogen peroxide[39b]; **cyclohexenyl hydroperoxide**, prepared by low temperature oxidation of cyclohexene to ≤ 21% conversion (74°, 36 p.s.i.g. O_2), **acetone**, Shell Chemical Co., redistilled from $KMnO_4$ and then from Na_2CO_3; **diethyl ketone, methyl *sec*-butyl ketone, pinacolone and biacetyl**, Eastman Kodak Co., white label, redistilled; **diobiacetyl**, Eastman Kodak Co., white label, redistilled; **dioxane**, commercial grade purified by standard procedure[40]; **cyclohexanone**, Eastman Kodak Co.; **2-chlorocyclohexanone**, Aldrich Chemical Co., distilled product, b.p. 71° (7 mm.), $n^{20}\text{D}$ 1.4833–1.4838; **2-methyl cyclohexanone**, Aldrich Chemical Co., G.L.C. showed one minor impurity; **2,2-dimethoxypropane**, Dow Chemical Co.

Solvents.—The alcohols ethanol, isopropyl alcohol, *n*-butyl alcohol and *t*-butyl alcohol were commercial chemicals which were redistilled after refluxing with calcium hydride. Others were reagent grade used directly.

***t*-Butyl Hydroperoxide with Metal Ions.**—Reactions were run in evacuated 150-ml. flasks fitted with dropping funnels. The ferrous solution was added to the stirred metal ion–hydroperoxide solution at temperatures which were kept constant with a water-bath. Analyses of the gaseous products were made directly on samples extracted from the flask and submitted to G.L.C. They were checked with mass spectra results. Methane 2′ and ethane 12′, were determined on a silica gel column at 0°; methyl chloride 8′ was determined on a dimethyl sulfolane column at room tem-

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[34] However, the absence of good model systems, heretofore, for the study of the reactions of free primary carbonium ions in solution, in part, vitiates this conclusion. The deamination reaction is often a good model system for the production of carbonium ions,[30] especially tertiary and allylic ones, as are the solvolytic reactions of halides and esters. The primary alkyl amines may not yield primary carbonium ions, but alkyl diazonium ions may be the true intermediates.[31b] In this case information regarding the reactions of free primary carbonium ions is not available and the *products from the oxidation of primary radicals cannot be categorically used to eliminate the carbonium ion mechanism.*

[35] L. Pauling, "The Nature of the Chemical Bond," 3rd Ed., Cornell University Press, Ithaca, N. Y., 1960, p. 57.

[36] The facile ligand transfer reactions between methyl radicals and cupric sulfate or chloride, for example, cannot involve participation by a β-carbon–hydrogen bond. The reaction of neopentyl radicals with cupric chloride to form neopentyl chloride does not seem to be adversely affected by the absence of β-hydrogens. Since the chlorine ligand transfer reaction in general is faster than the electron transfer reaction (elimination) it is reasonable to assume that the β-carbon–hydrogen bond does not play an important role in either reaction.

[37] N. Milas and D. Surgenor, *J. Am. Chem. Soc.*, **68**, 643 (1946).

[38] H. Williams and H. Mosher, *ibid.*, **76**, 2984 (1954).

[39] (a) N. Milas and L. H. Perry, *ibid.*, **68**, 1939 (1946); (b) N. Milas, R. L. Peeler and O. L. Mageli, *ibid.*, **76**, 2322 (1954).

[40] L. F. Fieser, "Experiments in Organic Chemistry," 2nd Ed., D. C. Heath Co., New York, N. Y., 1941.

perature. The aqueous solutions were analyzed qualitatively by G.L.C. using a triethylene glycol column at 100°. Acetone 6′, *t*-butyl alcohol 9.8′ and methanol 11.5′ could be determined easily even in dilute solutions containing metal ions. Prolonged usage of the columns with aqueous solutions led to bleeding of the absorbant. The crude reaction mixtures were also directly submitted for functional group analyses of total carbonyl and total hydroxyl values. The carbonyl value was determined by aqueous distillation into hydroxylamine hydrochloride solution for aldehydes and ketones and the total hydroxyl value used for the alcohols was obtained by the nitrite ester method.

Ferrous ion analysis was made by direct titration with ceric sulfate in sulfuric acid solution using a ferrous phenanthroline indicator. With highly colored copper solutions, the titrations were done potentiometrically with platinum and calomel electrodes.

***t*-Amyl Hydroperoxide and Metal Ions.**—These reactions were conducted in a manner similar to the experiments with *t*-butyl hydroperoxide. The gases were immediately analyzed by G.L.C. using the same silica gel column at 0°: CH_4, 2′; C_2H_6, 12′; C_2H_4, 18′. The aqueous solutions were analyzed with a triethylene glycol column: acetone, 6′; methyl ethyl ketone, 8.1′; ethanol, 15.5′; *t*-amyl alcohol, 20.5′.

Triptyl Hydroperoxide. Synthesis.—A solution of 174 g. (1.5 moles) of methyl pivalate in 500 ml. of ether was added to a solution of methylmagnesium chloride prepared from 175 g. (3.5 moles) of methyl chloride and 74 g. (3 moles) of magnesium turnings.[41] The reaction was worked up with aqueous ammonium sulfate and gave 86 g. (74%) of triptyl alcohol, b.p. 131–134°. A similar reaction of *t*-butyl Grignard reagent and acetone, though more convenient, gave a product which was contaminated with mesityl oxide in poor yield.

Triptyl alcohol was dehydrated by dissolving it (50 g.) in 50 ml. of methanol and adding the solution dropwise to a 50% aqueous sulfuric acid solution at 135°. The distillate was collected, washed with 10% sodium bicarbonate solution and redistilled; b.p. 78–80°, $n^{20}D$ 1.3995, yield 37 g. (80%) of 2,3,3-trimethylpentene-1.

Triptyl alcohol (10 g., 0.09 mole) was added dropwise at 0° to a solution of 15 g. of concd. sulfuric acid, 8 g. of water and $\sim$20 g. of 30% H_2O_2 at 10°. The reaction was stirred at 0° for 4 hours, then quenched in 10% sodium bicarbonate and extracted with ether. The ether extract was dried with sodium sulfate and distilled. A wide boiling mixture was obtained which consisted of acetone, *t*-butyl alcohol and triptyl alcohol. The hydroperoxide could not be distilled apart from the alcohol adulterant.

Triptene (2,3,3-trimethylpentene-1, 40 g., 0.41 mole) in 25 ml. of ether was stirred at 0° with a solution of 140 g. of hydrogen peroxide (30%) and 20 ml. of concd. sulfuric acid. The ether layer was extracted with 30% sodium hydroxide solution to precipitate the sodium salt of the hydroperoxide. The solid was filtered, washed with ether and redissolved in water. Triptyl hydroperoxide was sprung with carbon dioxide and the product extracted with ether. The ethereal solution yielded 5 g. (9%) of triptyl hydroperoxide. Variations in the sulfuric acid and hydrogen peroxide concentrations and temperatures were not very effective in increasing the yield.

Crystalline triptyl hydroperoxide is most conveniently prepared from the autoxidation of triptane at 118° and 35 p.s.i.g. oxygen. At 25% conversion a titrated yield of 46% hydroperoxide was indicated. The product was isolated by distilling the unreacted triptane *in vacuo*. The residual oil on further distillation resolved into *t*-butyl alcohol and essentially pure triptyl hydroperoxide. Further purification is effected by vacuum sublimation to produce a colorless solid smelling strongly of hydroperoxide; m.p. 26°, yield 23%, hydroperoxide equivalent 144.

Anal. Calcd. for $C_7H_{16}O_2$: C, 63.6; H, 12.2. Found: C, 63.5; H, 12.1.

Reaction with Metal Ions.—Triptyl hydroperoxide undergoes facile rearrangement in the presence of acids to acetone and *t*-butyl alcohol. In the presence of metal ion, this ionic process competes with the oxidation–reduction reaction. The exact amount of reaction proceeding *via* the latter process is, thus, difficult to ascertain. The change in ferrous titer, as indicated previously, is not a reliable index since ferrous ion is reformed from the reaction of *t*-butyl free radical and ferric ion. The yield of oxidation product, isobutylene, also is not an adequate basis for evaluating the oxidation–reduction reaction because an indeterminate number of *t*-butyl carbonium ions are solvated to *t*-butyl alcohol, a product of the ionic process as well. The yield of gaseous products formed was arbitrarily used as a crude estimate of the amount of oxidation–reduction reaction. These gaseous products consisted wholly of isobutane and isobutylene; they were resolved cleanly on a 20′ oxydipropionitrile–firebrick column at 0°: isobutane, 6.3′; isobutylene, 10.3′.

Photolysis of Ketones.—The apparatus[41a] as illustrated in Fig. 1 consisted of a Pyrex cell in which a water-jacketed Vycor thimble was immersed. The two outlet tubes were connected to (a) a manometer and (b) a bulb to collect gases. The system was freed of oxygen by successive evacuations to 1.5 cm. and flushing with oxygen-free nitrogen. The solution was stirred magnetically and the temperature controlled by regulating the amount of water passing through the water jacket. The light source consisted of the unfiltered radiation from the filament of a General Electric AH-4 mercury lamp. Gas yields were determined by the change in pressure in the system whose volume had been calibrated. Analyses were made by the G.L.C. described previously. In the most part, changes in metal ion concentrations were measured by analyzing the amount of precipitated copper. These values were checked by analyzing the solutions before and after irradiation for cupric ion electrodeposition. No attempt was made to measure quantum yields.

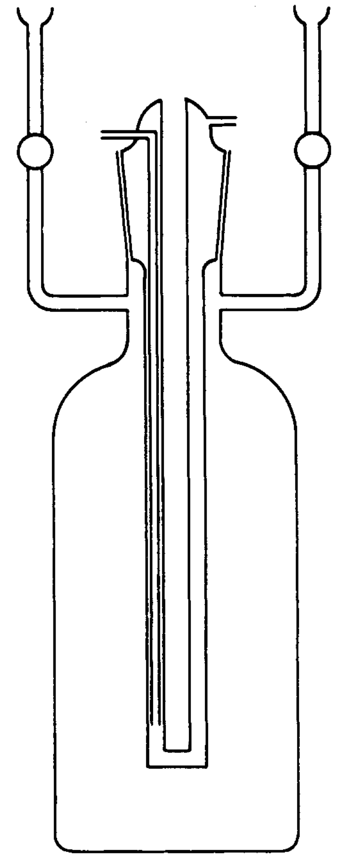

Fig. 1.—Photolysis cell.

Acetone and diethyl ketone photolyses were conducted in dilute aqueous solutions. The less soluble methyl *sec*-butyl ketone and pinacolone were photolyzed in 30% v. dioxane–water. Pure dioxane does not photolyze but does serve as a hydrogen donor for alkyl radicals.

The isomeric *n*-butenes were analyzed by G.L.C. in a 50′ dimethyl sulfolane column at room temperature: C_2H_6, 5′; C_2H_4, 5.6′; n-C_4H_{10}, 7.4′; 1-C_4H_8, 10.6′; *trans*-2-C_4H_8, 12.6′; *cis*-2-C_4H_8, 14.8′.

***n*-Butyl Hydroperoxide.**—The metal ion reactions were run in the conventional manner. The aqueous mother liquors were examined by vacuum distillation to dryness. The distillate was extracted once with ether. The ethereal solution of *n*-butyraldehyde was distilled to remove the ether and the residue treated directly with 2,4-dinitrophenylhydrazine. Recrystallizaton once from ethanol gave material, m.p. 116° (lit.[42] 122°).

Anal. Calcd. for $C_{10}H_{12}N_4O_4$: C, 47.6; H, 4.8; N, 22.2. Found: C, 46.9; H, 4.8; N, 21.8.

The aqueous distillate was also treated with Brady reagent to give a crystalline 2,4-dinitrophenylhydrazone, m.p. 164–165° (lit.[42] 166°).

Anal. Calcd. for $C_7H_6N_4O_4$: C, 40.0; H, 2.9; N, 30.5. Found: C, 40.9; H, 2.7; N, 31.1.

General Procedure for Reactions with "Cyclohexanone Peroxide."—Hydrogen peroxide (30%) was added to the cyclohexanone in a 500-ml. to 1-liter Morton-type 3-neck flask equipped with stirrer, thermometer and addition funnel. The flask and

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[41] *Cf.* E. Coburn, "Organic Syntheses," Coll. Vol. III, John Wiley and Sons, Inc., New York, N. Y., p. 696.

[41a] We wish to thank Dr. J. Kumamoto for the design of this cell.

[42] R. Shriner, R. Fuson and D. Curtin, "Systematic Identification of Organic Compounds," John Wiley and Sons, Inc., New York, N. Y., 1956.

stirrer were thoroughly cleaned and then washed with freshly prepared magnesium silicate suspension prior to each experiment. Addition of peroxide was made rapidly with the standard run (0.2 mole ketone: 0.1 mole H_2O_2), the temperature rising to ~35° with no external cooling.

After allowing 20–40 min. for formation of the cyclohexanone peroxide,[43] the metal ion solution was placed in the addition funnel and the system preswept with nitrogen. The reaction flask was cooled to, *e.g.*, ~20° and metal ion addition begun. Addition was conducted with vigorous stirring and at a rate which maintained temperature limits (~25 ± 5°) with the reaction kettle immersed in the ice-bath. For reactions run at 0° to −30°, Dry Ice–solvent-baths were employed. After addition of the metal ion solution, the reaction mixture was vigorously stirred for 1–2 hr. at 25 ± 5°. In reactions run at lower temperatures with cupric–ferrous sulfate, longer reaction times were employed. In these latter runs in the absence of organic solvent, disappearance of the solid peroxide adduct was used as a criterion of reaction time. Once the solid adduct had been largely consumed, the mixture was gradually warmed to room temperature.

Crude acid product was thoroughly extracted with chloroform (*ca.* four 75-ml. extracts) directly from the reaction mixture or from the mixture diluted with dilute sulfuric acid. The latter technique was necessary only where the *p*H was such that basic metal salts had precipitated during the reaction. The chloroform extract was dried over $MgSO_4$.

In runs made in aqueous methanol, work-up was carried out by diluting with water, and extracting with diethyl ether or chloroform. The extract was dried and concentrated by distillation. In either case the concentrates were usually fractionated in a micro-Vigreux distillation assembly.

Cyclohexanone Peroxide and Cupric–Ferrous Sulfate.—The chloroform extract resulting from the reaction of the subject peroxide (0.2 mole H_2O_2:0.4 mole ketone) with cupric–ferrous sulfate solution (100 g. of $CuSO_4 \cdot 5H_2O$, 0.4 mole; 111 g. of $FeSO_4 \cdot 7H_2O$, 0.4 mole; 22 ml. of concd. H_2SO_4 in ~400 ml. of water) was distilled to remove the bulk of the chloroform. The residue was fractionated in a micro-Vigreux column. After removal of unreacted cyclohexanone (fraction 1, 0.235 mole, b.p. 46–47° (23 mm.), $n^{20.6}\mathrm{D}$ 1.4506), three principal fractions were collected. Fraction 2 (2.0 g., b.p. 89–106° (21 mm.), $n^{20}\mathrm{D}$ 1.4421) was analyzed as 77% w. 5-hexenoic acid and 23% w. cyclohexanone. Fraction 3 (6.7 g., b.p. 107–109° (21 mm.), $n^{20.6}\mathrm{D}$ 1.4339) and fraction 4 (2.2 g., b.p. 91–94° (8 mm.), $n^{20.6}\mathrm{D}$ 1.4339) were identical with authentic 5-hexenoic acid (lit. values: b.p. 95–95° (7 mm.), $n^{20}\mathrm{D}$ 1.4338)[44] as determined by comparison of infrared spectra. Combined fractions 2 to 4 represent a yield of 0.092 mole of 5-hexenoic acid (46% yield on H_2O_2; 57% yield on cyclohexanone). Data obtained at 0 to 5° indicate that this yield can be increased to ≥ 70% based on ketone.

The kettle residue (1.3 g., 0.006 equiv. acid) obtained from the distillation contained acid equivalent to ~4% yield based on cyclohexanone; part of this acid was probably 6-hydroxyhexanoic acid. The balance of this residue, presumably cyclohexanone condensate and/or ester, accounted for another ~4% of the ketone.

Continuous ether extraction (~90 hr.) of the aqueous phase gave a 1.5-g. residue (~11 meq. of acid prior to ether removal) which did not distil at 154° (19 mm.). After this heating, the residue was analyzed. Found: C, 60.9; H, 8.9; ester value 0.646 equiv./100 g.; acid value 0.26 equiv./100 g. or total of ~4 meq. The elemental analysis (equivalent to $C_{6.0}H_{11}O_{2.2}$) and functional group analyses suggested a mixture of hydroxy acid, polyester and/or lactone, some of the esterification occurring during work-up. Infrared spectral analysis was consistent with this picture.

The aqueous phase (811 g.) from the extraction was freed from ether and submitted for "wet carbon" analysis. Found: C, 0.09 (≈0.73 g. of carbon): ~0.01 mole calcd. as 6-hydroxyhexanoic acid, ~6% yield based on cyclohexanone. Thus ~80% of the cyclohexanone was accounted for. The remaining 20% must be largely lost in handling. Functional group analysis of the chloroform extract prior to the isolation work (charge −400 mmoles ketone: 200 mmoles H_2O_2) showed the material balance indicated in Table III.

Table III

Product analysis	Equiv.	Ketone equiv. (moles)
Total acid	0.119	0.119
Steam-volatile acid	.113	(.113)(~80% yield)
Unsaturation	.121	...
Recovered carbonyl	.263	.263
Water-soluble acid[a]	.021	.021 (~15% yield)
Cyclohexanone accounted for		Σ = 0.40 (~100%)

[a] By isolation + "wet carbon" analysis.

Cyclohexanone Peroxide with Ferrous Sulfate and Indium Chloride.—The chloroform extract from the reaction (20 ± 5°) of cyclohexanone peroxide (0.2 mole ketone–0.1 mole H_2O_2) and ferrous sulfate–indium chloride (0.1 mole–0.1 mole, 0.2 mole of H_2SO_4, 100 ml. of H_2O) was dried and distilled. After removal of solvent, the residue was distilled in a micro-Vigreux column. Fraction 1 (b.p. 54–57° (25 mm.), $n^{20}\mathrm{D}$ 1.4509, 10.1 g.) had physical properties in agreement with those of cyclohexanone (recovery 0.103 mole).

Fraction 2 (b.p. 28° (6 mm.) to 93° (1 mm.), $n^{20}\mathrm{D}$ 1.4631, 0.6 g., acid value 0.345 equiv./100 g.) was shown by infrared analysis to be a mixture of acid, cyclohexanone and some unsaturated product, probably conjugated.

Fraction 3 (b.p. 95–98° (1 mm.), $n^{20}\mathrm{D}$ 1.4586, 7.1 g.) gave an infrared spectrum which was compatible with 6-chlorohexanoic acid; yield (fr. 3 + acid from fr. 2) ≈ 49% based on hydrogen peroxide.

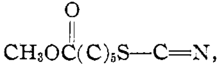
Anal. Calcd. for $C_6H_{11}O_2Cl$: C, 47.9; H, 7.36; Cl, 23.6; acid value, 0.665 equiv./100 g. Found: C, 48.9; H, 7.4; Cl, 23.3; acid value, 0.655 equiv./100 g.

Cyclohexanone Peroxide with Ferrous Chloride.—The ether extract from reaction of cyclohexanone peroxide (0.2 mole ketone: 0.1 mole H_2O_2) with ferrous chloride (0.2 mole; 0.2 mole H_2SO_4, 150 ml. of H_2O) at 0 to −5° was dried and distilled. After removal of the ether and recovered ketone (0.084 mole), 10.8 g. (~67% yield based on H_2O_2 and corr. for impurities) of crude 6-chlorohexanoic acid was obtained (b.p. ~109–120° (2 mm.), $n^{20}\mathrm{D}$ 1.4569–1.4596). However, this crude product turned pinkish-brown on standing in the light. Clearly, the crude chloro acid was contaminated with non-chlorine constituent presumably cyclohexanone and 5-hexenoic acid. Good agreement with the analyses was obtained on the basis of 3.8% and 2.7% of the latter components, respectively.

Anal. Calcd. for $C_6H_{11}O_2Cl$: C, 47.9; H, 7.36; Cl, 23.6; acidity, 0.665 equiv./100 g. Found: C, 49.8; H, 7.5; Cl, 22.1; acidity, 0.63–0.66 equiv./100 g.[45]; carbonyl value, 0.039 equiv./100 g.

Cyclohexanone Peroxide with Ferrous Thiocyanate.—The chloroform extract from a conventional reaction mixture (0.2 mole of ketone : 0.1 mole of H_2O_2; 0.2 mole of $FeSO_4$: 0.2 mole of NaSCN; 250 ml. of H_2O and 11 ml. of concd. sulfuric acid) gave a titratable yield of acid of 64% (based on H_2O_2). No solid acid production was observed. Distillation of the free acid indicated extensive product decomposition; therefore, conversion to the methyl ester was adopted as a preliminary step to distillation.

Methyl esterification was effected by (A) heating (~50°) a crude acid concentrate in the presence of excess methanol, 2,2-dimethoxypropane and Dowex-50 resin, or (B) by dissolving the cyclohexanone peroxide in acidic methanol-water (60:40 v./v.; 1 ml. of 6 *N* H_2SO_4), equilibrating for ~25 minutes at 25° and then adding ferrous sulfate and sodium thiocyanate.

Using procedure A, a product concentrate (0.057 mole of acid + ~0.05 mole of methyl ester from another reaction of 0.2 mole of hydrogen peroxide) was dissolved in 100 ml. of methanol containing 20 ml. of 2,2-dimethoxypropane. Dowex-50 resin (5 g.) was suspended in the methanol and the mixture stirred and heated at 40–60° until the bulk of the acid was gone. After filtering the catalyst and stripping off the methanol, the residue was distilled through a micro-Vigreux column. A yield of 14.0 g. (~0.075 mole) of product was collected at 121–124° (2 mm.), $n^{23}\mathrm{D}$ 1.4763–1.4756. The infrared spectrum was consistent with the postulated structure  $\mathrm{CH_3O\overset{\overset{O}{\|}}{C}(C)_5S{-}C{=}N}$, showing a strong $-C{\equiv}N$ band at 4.62 μ (shoulder at ~4.7–4.8) and a strong carbonyl band at 5.76 μ. *The product gave a negative color test for the isothiocyanate group.*[46]

Anal. Calcd. for $C_8H_{13}O_2SN$: C, 51.4; H, 7.0; N, 7.48; S, 17.1. Found: C, 52.1; H, 7.2; N, 7.13; S, 16.4.

2-Methylcyclohexanone Peroxide with Ferrous–Cupric Sulfate.—2-Methylcyclohexanone–hydrogen peroxide adduct (0.4:0.2 mole) was preformed (6 hr. equilibration with H^+) and reacted with cupric–ferrous sulfate (0.4:0.4 mole) at 20–40° to give ~54 % yield of total acid (*based on ketone*), ≥46% yield of steam-volatile acid. Distillation gave a 43% yield of acid (7.2 g., acid value 0.74 equiv./100 g. and bromine number 0.71 equiv./100 g.). Infrared analysis showed the presence of both "terminal" "*trans*" internal double bonds. Ozonolysis of distilled acid from two different runs and oxidation to acids were carried out. The *formate to total volatile acid ratios* indicated two types of un-

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[43] After ~5 min. from t_0 of mixing of ketone and peroxide, a homogeneous solution was obtained. If this solution were allowed to stand for longer periods of time, or if a trace of acid were added, a *white solid peroxide* would precipitate. In many reactions described here, the white solid precipitated prior to introduction of metal ions. Where this was not the case, one observed precipitation of the solid with the first addition of metal ion solution.

[44] M. Ansell and S. Brown, *J. Chem. Soc.*, 1788 (1957).

[45] Electrometric value 0.63; indicator value 0.66 equiv./100 g.

[46] E. Lieber and J. Ramachandran, *Chem. Ind.* (London), 1406 (1958).

saturation : *R*-143, 55% terminal : 45% internal; *R*-128,[47] 57–71% terminal:43–29% internal. These determinations assumed that the over-all production and isolation of formic and acetic acid were of equal efficiency in the ozonolysis–oxidation treatment.

The chloroform extract from another large scale reaction of 2-methylcyclohexanone peroxide (1:0.5 mole of ketone:peroxide)[48] with cupric–ferrous sulfate (0.5:0.5 mole; 300 ml. of H_2O, 0.2 mole of H_2SO_4) was distilled. After recovery of 0.67 mole of 2-methylcyclohexanone (b.p. 71–73° (31 mm.), n^{20}D 1.4482), 19.4 g. of crude acid (b.p. 63° (11 mm.) to 96° (5 mm.), n^{20}D 1.4478) was distilled;. acid value of 0.46 equiv./100 g., 18% yield based on hydrogen peroxide. The crude acid was contaminated with principally 2-methyl-cyclohexanone ($\leq$ 41% w., 0.07 mole). A viscous tarry kettle residue (4.7 g.) was not further investigated. An aliquot of the distilled acid was treated with ozone at −20 to −30° in methylene chloride. The ozonolysis products were oxidized by slight modification of Henne's hydrogen peroxide technique.[49] Distribution of the isomeric acids as to type of double bond (internal or terminal) was made by determination of total formate and total volatile acid as described (*vide supra*). Total "volatile acid" was obtained by vacuum stripping with *m*-xylene, a known aliquot of the aqueous oxidation mixture; duplicate analyses were in good agreement.

For infrared measurements, a 7.275-g. aliquot of crude distilled acid was "topped *in vacuo*" to remove 2.23 g. of unreacted methylcyclohexanone. This gave 5.043 g. of acid concentrate (acid value of 0.618 equiv./100 g.; ~79% w. as heptenoic acids) and reduced to a negligible amount the interference by methylcyclohexanone in the infrared analyses. Known concentrations of this "unknown acid concentrate" were prepared in chloroform and concentrations of "terminal" and "internal" double bond were determined by infrared analyses using the 10.05 and 10.34 peaks, respectively; found, ~60:40 ratio. The reference "standards" used for these measurements were 6-heptenoic acid and "*trans*-(?)"-2-methyl-3-hexenoic acid[50]; known concentrations of these standards were prepared in chloroform and used for calibration. Unfortunately, these infrared measurements accounted for only 70 to 75% of the acid content of the unknown. This discrepancy indicated that (a) some saturated acid was present, (b) the extinction coefficient of our models differed significantly from one or more of our unknowns, or (c) some "*cis*" internal acid was present. The fairly good agreement of infrared measurements with the ozonolysis data indicate, however, that this lack of closure is not serious. If, *e.g.*, one attributes the lack of closure to (c), then from infrared data one would calculate ~55% "internal" and ~45% "terminal" acid which differs only ~10% from the ozonolysis work.

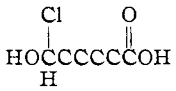
In an effort to determine the amount of 2-methyl-5-hexenoic acid, an aliquot of 4.145 g. was dissolved in 5% sodium hydroxide, treated with Norit decolorizing carbon, filtered, and the solution thoroughly extracted with ether. Regeneration of the acid, ether extraction and removal of volatiles *in vacuo* gave a residue of 2.921 g. (n^{20}D 1.4432, acid value 0.75 equiv./100 g.). This material was examined by n.m.r.[51] and the spectrum compared with the model compound, 2-methyl-3-hexanoic acid. It was concluded that <5% mole of the acid mixture was 2-methyl-5-hexenoic acid.

2-Chlorocyclohexanone Peroxide with Ferrous–Cupric Sulfate.—The reaction of 2-chlorocyclohexanone peroxide (0.4:0.2 mole of ketone:peroxide) with cupric–ferrous sulfate (0.2:0.2 mole) at 20 ± 5° was carried out in the conventional manner. Distillation of the chloroform extract gave a recovery of ~0.17 mole of 2-chlorocyclohexanone (b.p. ~89° (17 mm.), n^{20}D 1.4820–1.4889); infrared spectrum showed some contamination with conjugated impurity (probably aromatic) and hydroxyl-containing material.

The yield of distilled chloro acid was only 3.6 g. (b.p. ~101–109° (1–2 (mm.), n^{20}D 1.4682–1.4640). The low bromine number indicated the possibility of the presence of saturated acid as well as the failure of  $\mathrm{ClC{=}CC_3\overset{\overset{O}{\|}}{C}OH}$ to respond quantitatively to the method. The infrared spectrum was compatible with the unsaturated chloro acid structure.

Anal. Calcd. for $C_6H_9O_2Cl$: C, 48.6; H, 6.1; Cl, 23.9; acid value, 0.674 equiv./100 g. Found: C, 48.5; H, 6.6; Cl, 24.0; "stronger" carboxylic acid, 0.14 equiv./100 g.; "weaker" carboxylic acid, 0.52 equiv./100 g.; Br no. (by drastic method), 0.36 mole/100 g.

A large tarry kettle residue was obtained (7.3 g., acid value 0.52 equiv./100 g.). *Anal.* Found: C, 54.1; H, 6.7. Infrared spectrum showed a strong carbonyl, carboxylic hydroxyl and little or no C=C absorption. In another experiment (starting with freshly distilled chloro ketone) a similar result was obtained; analysis of the kettle residue gave 10.7% Cl. Therefore, this product was not simply C_6-aldehyde-acid expected from the possible intermediate



$$\mathrm{HO\underset{H}{\overset{\overset{Cl}{|}}{C}}CCCCC\overset{\overset{O}{\|}}{C}OH}$$

2-Chlorocyclohexanone with Ferrous Sulfate–Cupric Chloride.—The subject reaction was carried out as with the metal sulfates (*vide supra*), but with cupric chloride in place of cupric sulfate. Distillation of the chloroform extract gave 27.3 g. (b.p. ~85° (13 mm.), n^{20}D 1.4814–1.4840, 0.20 mole), of recovered 2-chlorocyclohexanone. Fraction 1 (b.p. 73–138° (3–4 mm.), 3.3 g.) had an acid value of 0.45 equiv./100 g. and fractions 2′–3′ (141–145° (3–4 mm.), 10.2 g.) had an acid value of 0.55 equiv./100 g. These chloro acids (1′ and 2′–3′) were examined by n.m.r. spectroscopy and an approximate assignment of composition was made

Fraction	$\mathrm{Cl_2CCCCCC\overset{\overset{O}{\|}}{C}OH}$	$\mathrm{ClCCCC\overset{\overset{Cl}{|}}{C}{-}\overset{\overset{O}{\|}}{C}OH}$	2-Chlorocyclohexanone
1′	~75% mole	~12.5% mole	~12.5% mole
2′–3′	~75% mole	~25% mole	

Potentiometric titration of the acids gave values in good agreement with these results

Fraction	$\mathrm{Cl_2CCCCCC\overset{\overset{O}{\|}}{C}OH}$ (pK_A ~4.8)	$\mathrm{ClCCCCCC\overset{\overset{O}{\|}}{C}OH}$ (Cl on C-2) (pK_A ~2.8)
1′	~0.44 equiv./100 g.[52]	~0.06 equiv./100 g.[52]
2′–3′	0.40 equiv./100 g.	0.13 equiv./100 g.

The infrared spectrum was compatible with the saturated chloro acid structures. A strong band at 13.35 μ may be due to "C–Cl stretch."

Anal. (fr. 2′–3′) Calcd. for $C_6H_{10}O_2Cl_2$: C, 39.0; H, 5.45; Cl, 38.3; acid value, 0.54 equiv./100 g. Found: C, 39.2; H, 5.5; Cl, 37.3.

Wet carbon analysis of the aqueous phase indicated a maximum "carbon" content of 0.36% w. ≈ 2.23 g. of carbon. Continuous extraction of the aqueous phase with ether permitted the isolation of ~1.7 g. of a semicrystalline residue. This crude residue gave a positive Schiff test, but an attempt to purify it by recrystallization was unsuccessful.

Synthesis of Model Compounds for Identification Purposes. 5-Hexenoic Acid.—Tetrahydrofurfuryl alcohol was converted to tetrahydrofurfuryl chloride and the latter cleaved using sodium sand in ether. The ring cleavage product, 4-penten-1-ol, was converted to 1-bromo-4-pentene by the action of phosphorus tribromide. The unsaturated bromide was then converted to the Grignard reagent and carbonated to give 5-hexenoic acid, b.p. 93° (9 mm.), n^{20}D 1.4337, acid value 0.885 equiv./100 g. (theory 0.876 equiv./100 g.; lit.[44] values: b.p. 95–97° (7 mm.), n^{20}D 1.4338).

6-Heptenoic Acid.—1-Bromo-4-pentene (described previously) was treated with sodium diethyl malonate, hydrolyzed and decarboxylated by conventional methods to give 6-heptenoic acid,[53] b.p. 127–128° (21 mm.), n^{20}D 1.4383, acid value 0.795 equiv./100 g. (theory 0.781 equiv./100 g.; lit. values: b.p. 125° (15 mm.), n^{15}D 1.4404).

2-Methyl-3-hexenoic Acid.—The acid, obtained through the courtesy of G. B. Payne, was distilled prior to use; heart cut had b.p. 120–121° (22 mm.), n^{20}D 1.4398, acid value 0.808 equiv./100 g. (theory 0.78 equiv./100 g.; lit. values: b.p. 118–120° (23 mm.), n^{20}D 1.4398[54]; 121–122° (24 mm.), n^{25}D 1.4382.[55] Although this acid was clearly principally "*trans*," the absence of "*cis*" isomer was not established.

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(47) The higher terminal value was obtained from the ratio of moles formate : moles organic acid reacted with ozone; the lower value from the ratio of moles formate produced : total moles of volatile acid. With run 143 these values were about the same.

(48) Hydrogen peroxide and the ketone were stirred or ~15 hours at room temperature and then for 2 hours in the presence of 1 ml. of 6 *N* sulfuric acid prior to introduction of metal ions.

(49) A. Henne, *J. Am. Chem. Soc.*, **65**, 753 (1943). The bulk of the unreacted hydrogen peroxide was destroyed with platinum black prior to analytical work.

(50) Authentic sample believed to be mostly "*trans*" was kindly furnished by Dr. G. B. Payne; see *J. Org. Chem.*, **24**, 1830 (1959).

(51) The authors are indebted to Dr. C. A. Reilly of this Laboratory for these measurements.

(52) Analysis obscured by some hydrolysis *in situ*; Σ value higher than that found with indicator method. The compound undergoing hydrolysis is probably 2-chlorocyclohexanone.

(53) P. Gaubert, R. P. Linstead and H. N. Rydon, *J. Chem. Soc.*, 1971 (1937).

(54) G. B. Payne, *J. Org. Chem.*, **24**, 1830 (1959).

(55) A. C. Cope and C. M. Hofmann, *J. Am. Chem. Soc.*, **63**, 3456 (1941).

Methylcyclohexyl Hydroperoxide with Ferrous–Cupric Sulfate (Water).—Methylcyclohexyl hydroperoxide (11 g., 0.052 mole, b.p. 39–41° (~1 mm.), n^{20}D 1.4642–1.4654; ~62% w. hydroperoxide, balance principally methylcyclohexyl alcohol) was dissolved in 85 ml. of methanol and ferrous–cupric sulfate (0.1:0.1 mole; ~100 ml. of water) solution was added with vigorous stirring at a rate necessary to maintain 25 ± 5°, stirred for 2 hours at 25 ± 5° and let stand overnight. The reaction mixture was diluted to ~1 liter and extracted with ether. The remaining aqueous phase was acidified to render it homogeneous, and re-extracted with ether. The combined ether extracts were washed with dilute bicarbonate and water, and dried. After distillation of the bulk of the ether a 10.2-g. concentrate was obtained whose infrared spectrum and G.L.C. analysis were consistent with principally unsaturated ketone and methylcyclohexyl alcohol (~1.4:1.0) and some ether. Found: Br no., 0.34 mole/100 g.; total 0.034 mole of double bond or max. of 65% yield as a C_7-monounsaturated ketone. The product was distilled through an efficient spinning band column.

Fractions 1 and 2 (73–76° (55 mm.), n^{20}D 1.4262–1.4281, 2.9 g.) were consistent with unsaturated ketone as a principal component. The infrared spectrum showed characteristic olefin bands at 6.08, 10.08 and 10.96 μ and a strong carbonyl band at 5.83 μ; hydroxyl impurity was present. The combined fractions were converted to a 2,4-DNP; recrystallized m.p. 72.5–73.5°. A mixed m.p. with an authentic sample (75–76.5°) of 2-heptanone 2,4-DNP was depressed, 66.5–68.5°.

Anal. Calcd. for $C_{13}N_4O_4H_{16}$: C, 53.4; H, 5.5; N, 19.2. Found: C, 53.7; H, 5.5; N, 19.3.

Fraction 3 (b.p. 76–82° (55 mm.), n^{20}D 1.4449, 1–2 g.) was shown by infrared analysis to be a mixture of 6-hepten-2-one and methylcyclohexyl alcohol; n^{20}D indicated ~43 and 57%, respectively (G.L.C. indicated ~48:52%).

Fractions 4 and 5 (b.p. 80–82.5° 55 mm.), n^{20}D 1.4573–1.4589, 1.7 g.) contained principally methylcyclohexyl alcohol.

The kettle residue was dissolved in ethanol and examined by infrared—strong carbonyl band at 5.84 μ and a strong band at 8.24 μ. Carbonyl analysis indicated a total of 0.0049 equiv., an upper limit to the yield of 2,13-tetradecanedione.

Methylcyclohexyl Hydroperoxide with Ferrous–Cupric Sulfate (Methanol–Water).—Methylcyclohexyl hydroperoxide (15.5 g., 69.7% w., 0.083 mole) was dissolved in 100 ml. of methanol and a solution of cupric sulfate (0.17 mole in 125 ml. of water) was added. To this vigorously stirred suspension, ferrous sulfate solution (0.17 mole/70 ml. of water) was added as rapidly as the temperature (25 ± 5°) could be maintained. Conventional work-up and fractionation gave the following fractions

Fraction	B.p., °C. (mm.)	Wt., g.	n^{20}D	Unsaturation,	mole[56]
1–4	69–77 (56)	7.8	1.4244–1.4263	0.0593	≤94% theory
5	75–77 (56)	2.4	1.4318	0.0187	
6	77–82 (56)	2.0	1.4502		
7	70 (27)	3.5	1.4584	....	
8	70–93 (27)	0.6	1.4449	....	
Residue		2.8	Crude solid	....	

G.L.C. analysis and infrared analysis of fractions 1–4 indicated principally 6-hepten-2-one (Br_2 no. 85% of theory) with small amounts of ether, 2-heptanone and methylcyclohexyl alcohol. G.L.C. analysis of fractions 5 and 6 showed this material to be principally 6-hepten-2-one (Br_2 no. 48% of theory) and methylcyclohexyl alcohol. Fraction 7 was ~95% methylcyclohexyl alcohol, and fraction 8 was ~50% methylcyclohexyl alcohol and 50% of an unidentified carbonyl component (G.L.C. analyses).

The kettle residue was recrystallized from Skelly-B to give 0.6 g., m.p. 72–74° (after working on a porous plate), and 0.1 g. (crop II) of m.p. 74–75.3° (lit.[22] value for 2,13-tetradecanedione: m.p. 76–77°). Analysis of the mother liquors showed a carbonyl content of 0.006 equiv., not further investigated.

2-Tetrahydropyranyl Hydroperoxide.—2-Tetrahydropyranyl hydroperoxide (25.2 g. of 93% w., 0.20 mole) was added to a cupric–ferrous sulfate solution (0.24 mole Cu^{++}, 0.24 mole Fe^{++}, 250 ml. of water) at 25 ± 5°. After ~1 hr. stirring at 24°, the mixture was heated at ~50° for 3 hours. Sulfuric acid (20 ml. of 95%–50 ml. of water) was added and heating continued for another 3 hours at 50 ± 10°.

The reaction mixture was steam distilled *in vacuo* until no more organic product appeared in the distillate; this crude distillate was fractionated in the spinning band column until water was obtained at the head. A product–water azeotrope was collected (b.p. 82.5°, n^{20}D 1.3612) and combined with subsequent water-rich fractions. After saturating the combined distillate with anhydrous potassium carbonate, there was obtained 3.3 g. of organic product, n^{20}D 1.4124. The infrared spectrum and G.L.C. analysis were consistent with principally 3-buten-1-ol; bromine number, 1.05 moles/100 g. (theory 1.38), ≈ 72% w. as 3-buten-1-ol or 0.033 mole (18% yield).

A water-insoluble phase (7.4 g., n^{20}D 1.4517) was separated by one ether extraction of the metal salt solution remaining after steam distillation. Infrared analysis of this product showed it to be a mixture of hydroxyl and carbonyl-containing constituents. Functional group analyses on this product indicated a mixture of diol and formate and non-formate ester

0.32	equiv./100 g.	formate ester
.16–0.24	equiv./100 g.	non-formate ester
.48	equiv./100 g.	hydroxyl
.08	equiv./100 g.	acid[57]
~4%		water

Anal. Found: C, 59.1; H, 10.0; O, 30.9 (≈ $C_{5.0}H_{10}$-$O_{2.0}$).

These functional group analyses demonstrated the incomplete hydrolysis of the formate esters. Complete identification of the mixture was not attempted, but the principal components were probably 1,8-octanediol diformate and monoformate, 1,8-octanediol and δ-valerolactone (non-formate ester.) All of these components have limited water solubility and are compatible with the infrared analysis of the unknown.

Continuous ether extraction of the aqueous phase ("wet carbon" showed the presence of 5.1 g. of carbon) was carried out and the extract fractionated through a micro-Vigreux column. After removal of some ether, formic acid and some water, three relatively high-boiling fractions (4,5,6) were collected.

Fraction 4 (b.p. 101–140° (20 mm.), n^{20}D 1.4251, 1.2 g.) was analyzed by functional group: acid value, 0.66 equiv./100 g.; ester value, 0.57 equiv./100 g.; and positive formate test.

Fraction 5 (b.p. 150–154° (20 mm.), n^{20}D 1.4407, 1.4 g.) contained: acid, 0.49 equiv./100 g.; ester, 0.69 equiv./100 g.; total formate, 0.72 equiv./100 g. Correcting the formate for free acid (0.72–0.49) indicates that only ~33% of the ester is formate.

Fraction 6 (b.p. 168–190° (2 mm.), n^{20}D 1.4450, 2.1 g.) contained: acid, 0.46 equiv./100 g.; ester, 0.76 equiv./100 g.; total formate, 0.65 equiv./100 g. Correcting formate for free acid (0.65–0.46) indicates that ~25% of the total ester is formate. The kettle residue (1.1 g.) analyzed for $C_{5.0}H_{8.25}O_{2.34}$.

Fractions 4 to 6 all showed strong carbonyl and hydroxyl absorbance, and strong absorbance at 8 to 8.8 μ and ~ 11.0 μ. Although not conclusively identified, the higher boiling formate ester was probably 4-hydroxyl-1-butyl formate, and the non-formate ester was presumably δ-valerolactone or derivatives thereof.

Surprisingly, reverse addition of metal ion solution to peroxide gave no *water-insoluble phase*. However, there was no change in the amount of steam volatile product, *i.e.*, **3-buten-1-ol**. There was isolated 3.2 g. of the unsaturated alcohol, n^{20}D 1.4193, unsaturation 1.27 moles/100 g. (lit.[58] value, n^{25}D 1.4189). The high-boiling product extractable from the metal ion phase by continuous extraction was similar to that isolated previously (formate + non-formate ester), ~$C_{5.0}H_{8.9}O_{2.6}$.

Cyclohexenyl Hydroperoxide with Cupric–Ferrous Sulfate.—Cyclohexenyl hydroperoxide (74.3% w., 0.1 mole) was suspended in water (50 ml.) and cupric–ferrous sulfate (0.2 mole–0.2 mole, 230 ml. of water) was added to the vigorously stirred suspension. Addition was begun at 10° (flask immersed in ice-bath); however, addition of a few ml. of the metal ion solution raised the temperature to 35–40°. Following this initial temperature rise, the remainder of the metal ion solution was added rapidly at 10–20°. The reaction product was extracted with ether, dried and distilled. **Fraction 1** (b.p. 84–93° (60 mm.), n^{20}D 1.4867, 7.4 g.) was analyzed by G.L.C. and infrared analysis, and shown to be principally 2-cyclohexenone (~77% w.); carbonyl value, 0.81 equiv./100 g. (=78% w. 2-cyclohexenone). G.L.C. analysis also showed ~22% w. 2-cyclohexenol, ~1% unknown. Identity of the principal peaks was established by correlation with the emergence pattern of an authentic mixture (Aldrich Chemical Co.) and infrared spectral comparison.

G.L.C. analysis indicated that **fraction 2** (b.p. 63° (30 mm.), n^{20}D 1.4837, 0.6 g.) was similar to fraction 1, but with slightly greater 2-cyclohexenol content. Correction for the 2-cyclohexenone content (0.09 mole/100 g.) of the starting peroxide gave a yield of 2-cyclohexenone of ≤58% based on hydroperoxide.

Addition of *t*-Butoxy Radicals to Butadine (Urea–Water).—The reaction of *t*-butyl hydroperoxide (0.18 mole) with butadiene (1.85 moles) and cupric–ferrous sulfate (0.2:0.2 mole, 200 ml. of water, 100 g. of urea) at 0 to −5° with vigorous stirring gave a low yield of *t*-butoxy*butenols* (0.022 mole, ~12% yield on hydroperoxide). Infrared analysis indicated that the structure of the principal component was $(CH_3)_3COCC(OH)C{=}C$ (VII). The allylic isomer of VII may have been present in small quantity

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(56) Bromine numbers by two different methods gave identical results.

(57) The acid was probably produced to some extent during analysis. If the acid is assumed to be all formic, then the non-formate ester has its max. value of 0.24 equiv./100 g. This water-insoluble phase was eliminated by adding metal ions to the peroxide.

(58) E. D. Amstutz, *J. Org. Chem.*, **9**, 310 (1944).

but was not detected by low sensitivity G.L.C. A kettle residue (2.6 g., ≤ 0.018 mole, $\sim C_{8.0}H_{10.4}O_{3.4}$), accounted for the only other petroleum ether-extractable product.

After accounting for water-soluble steam-volatile products, *t*-butyl alcohol and methanol ($\Sigma \leq 0.017$ mole) and acetone, there was still ~0.13 mole of *t*-butyl hydroperoxide unaccounted for. Continuous ether extraction of the aqueous salt solution was then carried out. Two fractions were obtained from the extract, ~1.8 g. (n^{20}_{D} 1.4445, b.p. 50–76° (1 mm.), $\sim C_{8.0}H_{16}O_{1.9}$), and 8.3 g. of a viscous glassy residue. The latter analyzed for $C_{9.0}N_{18}N_{1.9}O_{2.1}$ and had an average molecular weight of 187 [theory for $(CH_3)_3COCC(NHCONH_2)C{=}C$ (VIII), $C_9H_{18}N_2O_2$, mol. wt. 186]. Although not rigidly identified, the elemental analysis, water solubility and infrared data suggested the incorporation of urea into the butadiene-containing product.

Addition of *t*-Butoxy Radicals to Butadiene (Methanol–Water) (A).—*t*-Butyl hydroperoxide (0.18 mole, 97.6% w.) and butadiene (~100 g.) were dissolved in methanol (100 ml.) contained in a three-necked Morton flask equipped with stirrer, addition funnel, thermometer and Dry Ice condenser. After a presweep with nitrogen, cupric–ferrous sulfate solution (0.2:0.2 mole in 260 ml. of water) was added at a rate which would maintain the temperature at approx. −2° (ice-salt-bath). After ~30 minutes at −2°, the reaction mixture was brought gradually to room temperature; unconverted butadiene was distilled into a Dry Ice trap. The reaction mixture was extracted with petroleum ether, dried and distilled through a micro-Vigreux column. Separation was poor and the distilled fractions (*sans* cut 1, $n^{19.7}_{\mathrm{D}}$ 1.4164, 2.7 g.[59]) were combined and fractionated in a spinning band column

Fraction	B.p., °C. (100 mm.)	n_{D}^{20}	Wt., g.
1′	80–90	1.4139	1.0
2′–3′	90	1.4142	4.6
4′	90–110	1.4282	1.6
5′	110–116	1.4283	0.4
6′	116–117	1.4280	1.2

The principal component in **fractions 1′–3′**, was assigned on the basis of infrared analysis. Fraction 3′ was shown to be devoid of "*trans*" double bands and to have strong bands at 10.08, 10.83 ($> C{=}CH_2$) and 11.37 μ. The structure 1-*t*-butoxy-2-methoxy-3-butene was tentatively assigned.

Anal. Calcd. for $C_9O_2H_{18}$: C, 68.4; H, 11.5. Found: C, 67.1; H, 11.5.

Chromatography of fractions 1′–3′ (hydroxyl impurity) over silica gel and redistillation gave an analytically pure sample, n^{20}_{D} 1.4145. Found: C, 68.6, 68.9; H, 11.4, 11.6.

Infrared analysis of **fraction 6′** showed a strong "*trans*" band at ~10.34 and the structure was tentatively assigned as 1-*t*-butoxy-4-methoxy-2-butene. In contrast to fractions 1′–3′, where the $>C{=}C<$ absorption occurred at 6.05 μ the $> C{=}C <$ (?) absorption in fraction 6′ occurrred at ~5.96 μ.

Anal. Calcd. for $C_9O_2H_{18}$: C, 68.4; H, 11.5. Found: C, 68.2; H, 11.4.

Micro hydrogenation of fractions 2′–3′ and 6′, respectively, were carried out over platinum until no further hydrogen uptake was observed. Found: ~103% and ≥96% of theory, respectively, for the proposed structure.

Proof of Structure of Methyl Ethers of *t*-Butoxybutenol (B).—In another experiment the methanol–water ratio was increased (300 ml. of CH_3OH:260 ml. of H_2O) for a charge of 0.156 mole of hydroperoxide, 0.2 mole of Cu^{++} and 0.2 mole of Fe^{++}. An aliquot (16.2 g.) of the crude product from this run was hydrogenated over Pt in ethanol (uptake of ~0.081 mole, 0.088 expected from Br number) until no further uptake was observed. Fractionation in the spinning band column (at 102 mm.) permitted the isolation of fractions 2–3 (b.p. 89–92.5°, n^{20}_{D} 1.4005–1.4013, 4.1 g.) fraction 4 (b.p. 92.5–96°, n^{20}_{D} 1.4020, 3.6 g.), fraction 5 (b.p. 96–101°, n^{20}_{D} 1.4016, 1.4 g.) and fraction 6 (b.p. 101–102°, n^{20}_{D} 1.4053, 3.1 g.).

A 7.01-g. aliquot of combined fractions 2–4 was mixed with 25 ml. of 50% H_2SO_4 and de-*t*-butylated by warming to 70°, then holding at 55° for several hours. The reaction mixture was diluted with water and steam distilled. The distillate was saturated with potassium carbonate and 6.6 g. of crude product was recovered and dried. Fractional distillation in a micro column gave 2.0 g. of a product, b.p. 92–93° (102 mm.), n^{20}_{D} 1.4122; G.L.C. analysis and infrared analysis indicated ~95% w. of a hydroxylic component plus traces of water, *t*-butyl alcohol and methanol. The product was dried over calcium hydride overnight, and redistilled to give a heart cut (b.p. 91° (95 mm.), n^{20}_{D} 1.4158) whose infrared spectrum and physical properties were those expected for 2-methoxy-1-butanol (lit. values: b.p. 147.2°, n^{25}_{D} 1.4138[60]). It was concluded that fractions 2–4 contained principally 1-*t*-butoxy-2-methoxybutane and, therefore, assignment of structure to the lower boiling isomer in fractions 1′–3′ (*vide supra*, section A) was correct.

Anal. Calcd. for $C_5H_{12}O_2$: C, 57.6; H, 11.6. Found: C, 57.3; 57.5; H, 11.6, 11.6.

An aliquot of **fractions 5 and 6** (hydrogenated product) was resolved by G.L.C. (DC-710 immobile phase) and the higher boiling component trapped in a liquid nitrogen trap, distilled and dried over CaH_2, n^{20}_{D} 1.4093; this sample (~1.2 ml.) was de-*t*-butylated with 48% H_2SO_4 and the product isolated in the manner described previously. Micro distillation gave a product (~0.2 ml.) having b.p. < 68° (25 mm.), n^{20}_{D} 1.4188; however, it was contaminated with water and possibly some *t*-butyl alcohol; lit.[61] values for 4-methoxy-1-butanol: b.p. 63–64° (7 mm.), n^{20}_{D} 1.4213. An authentic sample of 4-methoxy-1-butanol was made *via* a Williamson synthesis from the sodium salt of 1,4-butanediol (0.4 mole of diol, 0.1 mole of sodium) and methyl iodide (0.15 mole); b.p. 84° (16 mm.), n^{20}_{D} 1.4202. Comparison of the infrared spectrum of the authentic 4-methoxy-1-butanol with the spectrum of the cleavage product showed that they were identical in all major bands. Minor variations in the spectrum of the cleavage product were attributable to impurities known to be present from G.L.C. analysis (*t*-butyl alcohol, water, etc.). Therefore, the principal component (high boiling) in fractions 5–6 was 4-*t*-butoxy-1-methoxybutane and the assignment of structure to fraction 6′ (*vide supra*) was correct (see section A).

Hydroxy Radicals and *t*-Butyl Alcohol.—Hydrogen peroxide (22.3 g. of 31%, 0.2 mole) and cupric–ferrous sulfate (0.2 mole: 0.2 mole; 10.2 ml. of concd. H_2SO_4, 235 ml. of H_2O) were added simultaneously and with good stirring to a mixture of *t*-butyl alcohol (0.5 mole), sulfuric acid (4.5 ml.) and water (154 g.). Temperature was maintained at 22–25° and the mixture stirred for 30 minutes after completion of the addition. The crude mixture was analyzed for α-glycol; found 0.012 mole/100 g. = 36% yield based on peroxide.[62] The reaction mixture was concentrated *in vacuo* (50–70°), but the distillate contained isobutyraldehyde (Schiff test was positive; carbonyl value, 0.005 equiv./100 g. or total of 0.020 equiv.). Isobutyraldehyde was identified as its 2,4-DNP by a mixed chromatogram with an authentic sample; another carbonyl component (*via* mixed chromatogram test with the 2,4-DNP's) was indicated to be acetone and not formaldehyde. At any rate, production of isobutyraldehyde was consistent with the known reaction of isobutylene glycol in dilute acid solution. No further attempt was made to recover isobutylene glycol from the metal ion solution.

Analytical Methods.—The functional group analyses reported here were carried out for the most part by conventional methods of this Laboratory. Those requiring special comment are listed:

Total Formate.—Samples containing peroxide were subjected to ion exchange treatment prior to oxidation with $HgCl_2$.

α-Glycol.—The metal ion solution was passed through an ion exchange column prior to analysis for α-glycol.

Unsaturation values were determined usually by a "bromine number" method which was not interfered with by other functional groups (*e.g.*, $-\underset{\underset{\mathrm{O}}{\|}}{\mathrm{C}}-$) known to be present.

Infrared Analyses.—Quantitative work was done on the Perkin-Elmer 21, speed 3 to 4, suppression 2.

Ferrous ion was determined by titrating with standard ceric sulfate (ferrous phenanthroline indicator) or by potentiometric titration.

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(59) The major component of this material was shown to be identical with that found in fractions 1′–3′ (g.l.c. analysis).

(60) P. D. Bartlett and S. D. Ross, *J. Am. Chem. Soc.*, **70**, 926 (1948).

(61) R. H. Palomaa and R. Jansson, *Ber.*, **64B**, 1606 (1931).

(62) U. S. Patent 2,700,051 (to du Pont Co.) reports a yield of 0.175 mole of 2,5-dimethyl-2,5-hexanediol/1.06 moles of peroxide when ferrous sulfate is used alone.