S 3. *Some Oxidations involving the Free Hydroxyl Radical.*

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The free hydroxyl radical, produced by the action of ferrous salts on aqueous hydrogen peroxide, has been found to oxidise organic compounds of many types. A kinetic examination has shown that these oxidations fall into well-defined groups as (*a*) chain reactions or (*b*) non-chain reactions, and this grouping can be related to the structure of the free organic radical which is formed during the oxidation process. Certain of these organic free radicals are capable of reducing mercuric salts, or iodine.

Several substances, *e.g.*, acetic acid, acetone, chlorides, and bromides, which are not themselves oxidised under these conditions, have been found to retard the oxidations of other oxidisable organic substrates. This retarding action can be ascribed to a radical transfer which replaces the active hydroxyl radical by a much less active radical derived from the retarding agent. A preliminary kinetic study of this action indicates that characteristic " retardation coefficients " can probably be assigned to these radical-transfer agents.

THE work of Evans and his colleagues (*J.*, 1947, 266: *Trans. Faraday Soc.*, 1946, **42**, 155) has now confirmed the Haber–Weiss theory (*Proc. Roy. Soc.*, 1934, *A*, **147**, 332) that the reaction between ferrous ions and hydrogen peroxide occurs in two distinct stages :

$$(1)\ \mathrm{Fe^{++} + HO{-}OH \xrightarrow{k_1} Fe^{+++} + (HO\colon)^- + {\cdot}OH}$$

$$(2)\ \mathrm{Fe^{++} + {\cdot}OH \xrightarrow{k_2} Fe^{+++} + (HO\colon)^-}$$

Fenton's reaction—the oxidation of hydroxylic compounds by hydrogen peroxide in the presence of a ferrous salt—therefore involves free hydroxyl radicals, and, in an initial study of the oxidation of alcohols with Fenton's reagent, we have been able to show that the kinetics of this reaction can, by a suitable choice of conditions, be made relatively simple (compare Merz and Waters, Faraday Soc. Discussion, 1947, **2**, 179).

Our experimental study of oxidations involving free hydroxyl radicals which have been generated *in situ* by reaction (1) above has now been extended to include also the oxidations of water-soluble organic compounds of other types. Since to a large extent we have been able to elucidate the various mechanisms of these oxidations by adopting a kinetic approach, our results can best be described in terms of the following physicochemical theory.

Kinetic Features of Chain and Non-chain Oxidation by Hydrogen Peroxide.—If hydroxyl radicals, generated in dilute acid solution by process (1), are allowed to react with an organic compound, R–H, then there usually can occur the dehydrogenation process (3)

$$(3)\ \mathrm{HO{\cdot} + H{-}R \xrightarrow{k_3} H_2O + {\cdot}R}$$

which generates the free organic radical, •R. This radical may, in some cases, be able to react with molecular hydrogen peroxide, as in (4)

$$(4)\ \mathrm{R{\cdot} + HO{-}OH \xrightarrow{k_4} R{-}OH + {\cdot}OH}$$

by a process which involves an electron-transference somewhat similar to reaction (1), but, in other cases, the organic radical, •R, may be too stable to attack the hydrogen peroxide molecule.

If both reactions (3) and (4) can occur, then the induced reaction of R–H by hydrogen peroxide proceeds as a *chain reaction* in which a small amount of a ferrous salt may initiate the oxidation of a large amount of the organic substance. If, in contrast, reaction (4) cannot occur, then the organic radical •R may become stabilised in either of two ways: it may combine with a further hydroxyl radical, by a process (5) which is analogous to (2), or, alternatively, two radicals, •R, may meet and stabilise each other, either by dimerisation or by disproportionation (6).

$$\text{(5) } R\cdot + \cdot OH \xrightarrow{k_5} R\text{–}OH$$

$$\text{(6) } 2R\cdot \xrightarrow{k_6} \text{Inert products}$$

Processes (5) and (6) are kinetically indistinguishable in these circumstances. Chemical investigation of reaction products usually rules out dimerisation of radicals as a dominant chain-ending reaction, and only in certain cases does disproportionation of the organic radical appear to offer a plausible explanation of the formation of the observed reaction products.

Oxidations involving (5) or (6), and not (4), are *non-chain* processes in which at least half the hydrogen peroxide oxidises the ferrous salt and not the organic compound. The utilisation of more than 50% of the hydrogen peroxide by the organic compound indicates either the occurrence of reaction (4) or the intervention of consecutive reactions involving oxidation by the ferric ion, but there is available a still more decisive kinetic criterion for differentiating between the *chain* and the *non-chain* oxidation mechanisms.

If dilute hydrogen peroxide is added, with adequate stirring, to a mixture of an organic compound R–H and a ferrous salt, in such proportions that some ferrous salt remains in the system at the end of the reaction, then it has been found that the relative quantities of R–H and of Fe^{++} which are oxidised by the hydrogen peroxide are independent of the rate of addition of this oxidiser. As explained in our introductory paper (*loc. cit.*), the kinetic equations (A), (B), and (C) hold for a chain reaction comprising only reactions (1)—(4) above

$$\text{(A) } \frac{-d[\text{R–H}]}{-d[\text{Fe}^{++}]} = \frac{k_3[\text{R–H}]}{2k_2[\text{Fe}^{++}]} \quad \text{or} \quad \log\frac{[\text{R–H}]_t}{[\text{R–H}]_0} = \frac{k_3}{2k_2}\log\frac{[\text{Fe}^{++}]_t}{[\text{Fe}^{++}]_0}$$

$$\text{(B) } \frac{-d[\text{H}_2\text{O}_2]}{-d[\text{Fe}^{++}]} = 0{\cdot}5 + \frac{k_3[\text{R–H}]}{2k_2[\text{Fe}^{++}]}$$

$$\text{(C) } \frac{-d[\text{H}_2\text{O}_2]}{-d[\text{R–H}]} = 1 + \frac{k_2}{k_3}\frac{[\text{Fe}^{++}]}{[\text{R–H}]}$$

If, however, we are dealing with the *non-chain* sequence of reactions (1), (2), (3), and (5), but not (4), then the alternative equations are (E), (F), and (G):

$$\text{(E) } \frac{-d[\text{Fe}^{++}]}{-d[\text{R–H}]} = 2 + \frac{2k_2}{k_3}\frac{[\text{Fe}^{++}]}{[\text{R–H}]} \quad \text{[compare the inverse of (A)]}$$

$$\text{(F) } \frac{-d[\text{Fe}^{++}]}{-d[\text{H}_2\text{O}_2]} = 1 + \frac{k_2[\text{Fe}^{++}]}{(k_2[\text{Fe}^{++}] + 2k_2[\text{R–H}])}$$

$$\text{(G) } \frac{-d[\text{H}_2\text{O}_2]}{-d[\text{R–H}]} = 2 + \frac{k_2}{k_3}\frac{[\text{Fe}^{++}]}{[\text{R–H}]} \quad \text{(compare C)}$$

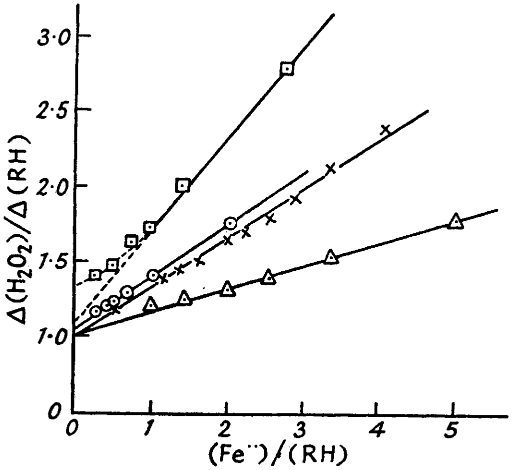
Use of equation (6) in place of (5) affects only the absolute value of the term k_2.

It will be seen, therefore, that the plotting of relative consumptions of hydrogen peroxide and of ferrous ion, according to the requirements of equations (C) and (G) will enable us to make a clear differentiation between chain and non-chain processes.

Figs. 1 and 2 give representative examples of substances which fall into the two classes. Naturally, the differential equations given above do not apply strictly, since minor side reactions, such as alternative chain-ending processes and consecutive oxidations of primary reaction products, have been neglected. The experimental work shows, however, that discrepancies become significant only when the ratio $[Fe^{++}]/[R\text{–}H]$ becomes small. Some of the implications of these deviations from equations (A), (C), and (G) are discussed below (pp. S 22, S 23). We would, moreover, at this stage, emphasise that the equations given above hold only for *acid* conditions in which the ionisation of molecular hydrogen peroxide is negligible.

It was shown by Haber and Weiss that other " catalase " reactions of hydrogen peroxide set in when anions, such as (˙O–OH)$^-$, and radical-ions, such as (·O–O˙)$^-$, may be present. For this reason we have regularly worked with dilute (M/50 to M/100) solutions in the presence of N/10-sulphuric acid; *i.e.*, at the constant pH = 1.

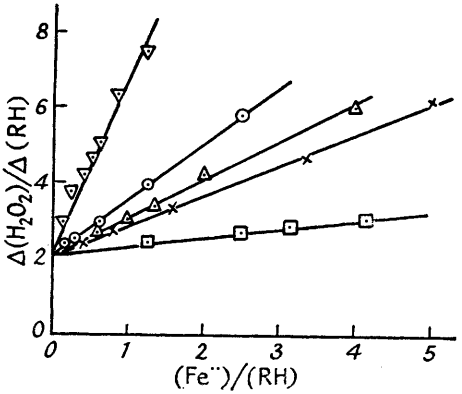
FIG. 1.



Substances oxidising by a chain reaction.

⊡ = *Lactic acid.* × = iso*Propyl alcohol.*
⊙ = *Formaldehyde.* △ = *Diethyl ether.*

FIG. 2.



Substances oxidising by a non-chain reaction.

▽ = *Pyridine.* × = *Pinacol.*
⊙ = tert.-*Butyl alcohol.* △ = *Methyl acetate.*
⊡ = *Chloral hydrate.*

Classification of Experimental Results.—Our results, which relate to measurements made at room temperature, show that water-soluble organic substances can be classified into three fairly well-defined groups, as shown in Tables I, II, and III.

The grouping of substances into categories (*a*) and (*b*) can quite clearly be associated with the chemical structure of the radical ·R, obtained by the first dehydrogenation process (3). It may be noted that all compounds of Table I form carbon, or possibly oxygen, radicals in which the electron-density at the point of free valency is relatively high. Thus in the simple primary or secondary alcohols, the ethers, and the aldehydes, the inductive (plus " hyper-

TABLE I.

Compounds which oxidise by a chain reaction.

Substance.	k_3/k_2.*
Primary and secondary alcohols:	
Methyl	2·3
Ethyl	3·8
n-Propyl	2·6
*iso*Propyl	3·0
1 : 3-Butylene glycol	(Multi-stage)
Hydroxy-acids:	
Glycollic acid	2·0
Lactic acid	1·7
α-Hydroxybutyric acid	3·0
(Thioglycollic acid	2·7)
Ethers:	
Diethyl ether	6·4
Dioxan	(Multi-stage)
Tetrahydrofuran	6·2
Tetrahydropyran	4·5
Aldehydes:	
Formaldehyde	3·0
Acetaldehyde	2·2
Amino-acids:†	
Glycine	0·035
Alanine	0·12

* k_3/k_2 has in each case been determined graphically from equation (C) for the reaction in N/50-solution in the presence of N/10-sulphuric acid.

In our introductory paper the figures were based upon equation (A) and calculated in equivalents and not moles, *i.e.*, upon the ratio $k_3/2k_2$, using a curve in which the extrapolation of the linear portion is often uncertain.

† These oxidations may not be simple chain processes. We have confirmed the statements of Wieland and Franke (*Annalen*, 1927, **457**, 1) that α-amino-acids promote some " catalase " decomposition of hydrogen peroxide.

TABLE II.

Compounds which oxidise by a non-chain reaction.

Substance.	k_3/k_2.	Substance.	k_3/k_2.
Tertiary alcohols:		*Carboxylic acids:*	
tert.-Butyl alcohol	0·65	Propionic acid	0·86
tert.-Amyl alcohol	—	*n*-Butyric acid	0·72
Pinacol	1·25	*iso*Valeric acid	3·3
Phenyldimethylcarbinol	—	Succinic acid	0·03
Esters:		Adipic acid	1·25
Methyl acetate	0·85	*Amines* (as their sulphates) :	
Ethyl acetate	1·1	Diethylamine	0·4
*iso*Propyl acetate	1·25	Triethylamine	0·8
Chloral hydrate	4·7	Pyridine	0·18

1 : 2-*Glycols*: *	(k_3/k_2 for persulphate oxidation).
Ethylene glycol	0·36
2 : 3-Butylene glycol	0·010
1 : 2-Butylene glycol	0·019

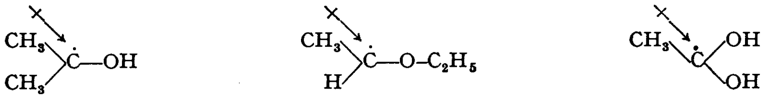
* The glycols were studied with persulphate and not hydrogen peroxide since in the latter case the consecutive oxidation of the aldehydic first products was extensive. For ethylene glycol the value of k_3/k_2 for H_2O_2 oxidation was about 4·0.

TABLE III.

Compounds not noticeably attacked by the hydroxyl radical.

Carboxylic acids :	Acetic acid, malonic acid, maleic acid, fumaric acid.
Ketones :	Acetone, methyl ethyl ketone.
Amides :	Urea, acetamide.

conjugation ") effect of the alkyl groups is such as to increase the electronegativity at the site of the free valency : thus



(*a*) *Molecules oxidised by the chain process.*

whilst in the carboxylic acids, or esters, or again chloral hydrate, there is present a strongly dipolar group of electron-abstracting character :

(*b*) *Molecules oxidised by the non-chain process.*

It is reasonable therefore to suppose that a free radical, •R, can only attack molecular hydrogen peroxide if it has an electron-availability (? reduction potential) over a certain value.

Properties and Structures of the Organic Free Radicals.—(1) *Reducing action.* Following up observations of Dhar (*J.*, 1917, **111**, 690; *J. Indian Chem. Soc.*, 1928, **5**, 203) and Weiss (Faraday Soc. Discussions, 1947, **2**, 188, 211) in a related field, we have found that those free organic radicals which can attack molecular hydrogen peroxide will also give up an electron to a mercuric ion, or to iodine, both of which they will reduce. For instance, when mercuric chloride is added to a mixture of a ferrous salt and an organic compound of group (*a*), *e.g.*, ethyl alcohol, there is no reaction, but the moment that hydrogen peroxide is added to this mixture then mercurous chloride is precipitated, since reaction (7) sets in, and competes with (4) in removing alcohol from the system. Mercuric nitrate acts similarly.

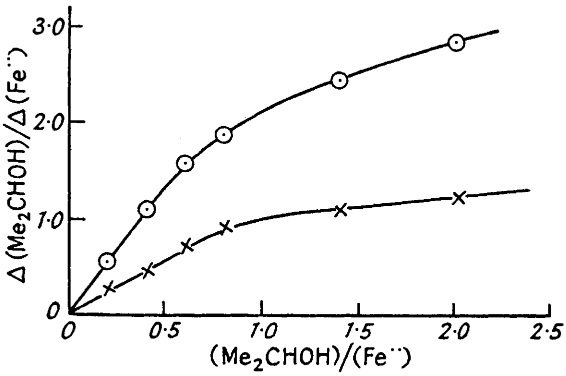
$$(7,a)\quad CH_3\text{-}\dot{C}H\text{-}OH + Hg^{++} \longrightarrow Hg^{+} + CH_3\text{-}\overset{+}{C}H\text{-}OH$$

$$(7,b)\quad CH_3\text{-}\overset{+}{C}H\text{-}OH \longrightarrow CH_3\text{-}CH{=}O + H^{+}$$

The reduction of the mercuric ions is not due to the production of an aldehyde: acetaldehyde and *iso*propyl alcohol both exhibit this effect extremely well when they are treated in this way. Since (7,*a*) competes with (4), the addition of the mercuric salt diminishes the fraction of hydrogen peroxide which is reduced by the organic substance, and consequently the ratio k_3/k_2 decreases in the way shown by Fig. 3. Iodine behaves similarly, and reaction (7,*c*) can be followed by adding an excess of iodine to the oxidisable mixture and, after running in the hydrogen peroxide, titrating rapidly with thiosulphate, since the Fe^{++}–I_2, the H_2O_2–I^- and the $CH_3{\cdot}CHO$–I_2 reactions are all comparatively slow.

$$(7,c)\ 2CH_3\text{–}\dot{C}H\text{–}OH + I_2 \longrightarrow 2I^- + 2CH_3\text{–}\overset{+}{C}H\text{–}OH$$

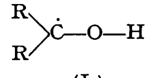
FIG. 3.

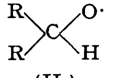


Effect of mercuric nitrate on the oxidation of iso*propyl alcohol.*
⊙ = *Oxidation without* Hg^{++}. × = *Oxidation with* Hg^{++}.
The mercuric nitrate was present in 0·05M-*concentration in the mixture.*

Very few of the active organic radicals which we have studied appear to be capable of reducing ferric ions rapidly in the same way, since the initial addition of an excess of ferric sulphate does not significantly affect the k_3/k_2 ratio. It may be noted that a reduction of Fe^{+++} by free radicals would increase the observed value of $\Delta[R\text{–}H]/\Delta[Fe^{++}]$ and this might have the effect of making an initial non-chain oxidation appear to proceed by a chain sequence.

Of the substances listed in Table II, only the ethylene and butylene glycols have any immediate reducing action upon mercuric salts in these circumstances. Essentially the oxidising or reducing properties of an organic free radical depend upon the environment of the unpaired electron. Thus, whilst radicals derived from primary or secondary alcohols by dehydrogenation can be formulated as either (I) or (II), tertiary alcohols can only yield free


(I.)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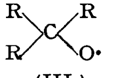
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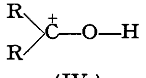
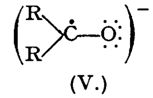
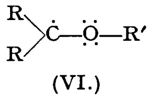

(IV.)

oxygen radicals (III). Free carbon radicals, *e.g.*, (I), can much more easily lose one further electron to give kations of known types [*e.g.*, (IV), the proton-adduct of a ketone] than gain an electron to form a free carbon anion, and hence are predominantly reducing agents. In contrast, the oxygen radicals (II) and (III) are oxidising agents since they tend to acquire but not to lose an electron. Both the reducing radical (I) and the oxidising radical (II) can, however, by loss of proton give the same mesomeric radical-anion (V) in which it is impossible to associate the odd electron exclusively with either the carbon or the oxygen atom. To a certain extent, too, this resonance of the other electron can occur in a radical of type (I), and


(V.)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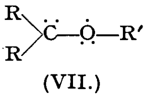

(VII.)

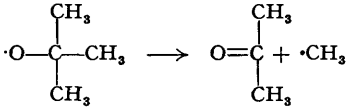
in the corresponding radicals derived from ethers or esters, since the two canonical structures (VI) and (VII) must both be taken into consideration. One would associate (VI) with reducing and (VII) with oxidising properties. Of these states (VI) will undoubtedly predominate, but nevertheless none of these radicals should be considered exclusively as reducing agents (see below, p. S 24).

(2) *The instability of radicals derived from tertiary alcohols.* The fact that *tert.*-butyl alcohol and its homologues are easily oxidised by free hydroxyl is noteworthy, since tertiary carbinols are remarkably resistant to many oxidising agents, such as acid dichromate solution, under conditions in which primary and secondary carbinols readily react. Again, *tert.*-butyl alcohol gives a remarkably stable hydroperoxide (Milas and Surgenor, *J. Amer. Chem. Soc.*, 1946, **68**, 205). Both acetone and formaldehyde were identified as products of the oxidation of *tert.*-butyl alcohol with Fenton's reagent. Evidently the non-reducing radical (III; R = Me) is first produced from the carbinol by a type (3) reaction, *viz.*

$$(8)\ \mathrm{HO\cdot} + \mathrm{H{-}O{-}C(CH_3)_3} \longrightarrow \mathrm{HO{-}H} + \mathrm{\cdot O{-}C(CH_3)_3}$$

and then breaks down either (i) by reaction with another hydroxyl radical (9) or (ii) by spontaneous fission to acetone and free methyl (10) which is thereupon oxidised by a non-chain reaction to methyl alcohol, and subsequently to formaldehyde :

$$(9)\ \mathrm{H{-}O\cdot} + \mathrm{(CH_3)_3C{-}O\cdot} \longrightarrow \mathrm{HO{-}CH_3} + \mathrm{(CH_3)_2C{=}O}$$

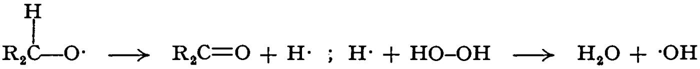
(10) 

$$\mathrm{\cdot O{-}C(CH_3)_2{-}CH_3} \longrightarrow \mathrm{O{=}C(CH_3)_2} + \mathrm{\cdot CH_3}$$

Under our experimental conditions we were unable to separate the small amount of methanol as such, but showed that an excess of oxidiser produced an appreciable amount of formaldehyde.

The recent demonstration by Raley, Rust, and Vaughan (*J. Amer. Chem. Soc.*, 1948, **70**, 88) and by Rust, Senbold, and Vaughan (*ibid.*, p. 95) that di-*tert.*-butyl peroxide vapour decomposes spontaneously at 130° or above into the free *tert.*-butyloxy-radical (III; R = Me) which then splits, in the gas phase, to acetone and free methyl supports reaction sequence (ii) (*i.e.*, 10) above.

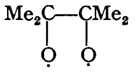
The oxidations of primary alcohols can be explained in a similar way provided (*a*) that it be assumed that free hydrogen is formed, and that this, unlike free methyl, can continue a *chain* reaction by attacking molecular hydrogen peroxide, and (*b*) that we discount the fact that a C–H bond has a slightly higher bond strength than a C–C bond : *e.g.*,

(11) 

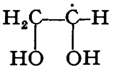
$$\mathrm{R_2C(H){-}O\cdot} \longrightarrow \mathrm{R_2C{=}O} + \mathrm{H\cdot}\ ;\ \mathrm{H\cdot} + \mathrm{HO{-}OH} \longrightarrow \mathrm{H_2O} + \mathrm{\cdot OH}$$

However, it is not possible to use mechanism (11) to explain the rapid oxidations of ethers by a chain reaction. This oxidation, which yields aldehydes, is noteworthy in that it accords so well with their facile reactions with free phenyl radicals produced from benzenediazoacetate (Waters, *J.*, 1937, 113), benzenediazonium chloride (Hanby and Waters, *J.*, 1939, 1793), and dibenzoyl peroxide (Cass, *J. Amer. Chem. Soc.*, 1947, **69**, 500).

The oxidation of pinacol to acetone is obviously a non-chain reaction analogous to the oxidation of *tert.*-butyl alcohol, and must proceed *via* the oxygen di-radical (VIII). The fact that Fenton's reagent can effect a similar C–C bond fission in other 1 : 2-glycols has already been noted by one of us (*Nature*, 1946, **158**, 380), but we have now shown that, in general,



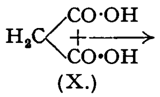
(VIII.)



(IX.)

this is no more than a minor side-reaction, although 1 : 2-glycols, unlike 1 : 3-glycols oxidise by the non-chain mechanism (see Fig. 4). Thus ethylene glycol yields mainly glyoxal and oxalic acid, whilst 2 : 3-butylene glycol yields mainly diacetyl. Hence the initial attack of the free hydroxyl radicals must be mainly, though not exclusively, on the C–H bonds of the carbinol so as to form the reducing radical of type (IX). The absence of a chain reaction from (IX) indicates that radicals of this type must be somewhat less reactive than the radicals derived from monohydric primary or secondary alcohols, and presumably this must be due to the proximity of the second hydroxyl group.

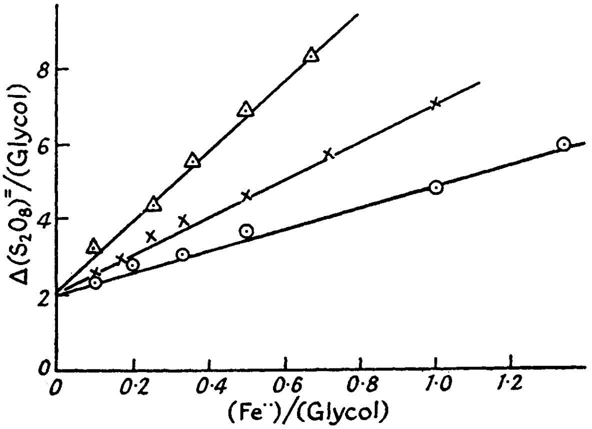
(3) *Other cases of non-chain reactions.* Acetic acid is not oxidised by hydroxyl radicals, but its homologues are undoubtedly attacked at a $-CH_2-$ group. Thus propionic acid yields some acetaldehyde, presumably *via* formation of lactic acid. It may be suggested that the hydrocarbon chains of these acids exert an electron-donating effect which is sufficiently powerful to overcome the electron-abstracting effect of the carboxyl group to a sufficient extent to weaken the C–H link for its dehydrogenation by free hydroxyl. In this connection the resistance towards oxidation of malonic acid (X) and the very slight oxidation of succinic acid are both significant. The fact that it is alkyloxy, R_2CH-O, and not the acyloxy, $R-CO-O-$, groups of esters which are attacked illustrates the importance of the spatial proximity of the polar carbonyl group in the un-ionised molecules. The inertness of the ketones may be ascribed to a similar origin, and evidently the inductive effects which favour proton release from a C–H link are the inverse of those favouring its oxidation (compare Waters, *J.*, 1946, 1153). The relatively low values of k_3/k_2 for the kations of amines also accord with this view.



$$H_2C\begin{matrix}\diagup CO{\cdot}OH \\ + \!\!\longrightarrow \\ \diagdown CO{\cdot}OH\end{matrix}$$

(X.)

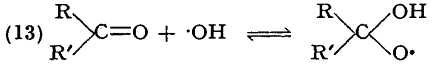
FIG. 4.



Oxidations of 1 : 2-*glycols with persulphate.*
⊙ = *Ethylene glycol.* △ = 2 : 3-*Butylene glycol.*
× = 1 : 2-*Butylene glycol.*

Radical-transfer in Reactions involving the Hydroxyl Radical.—It has already been surmised that equilibria leading to the formation of other oxygen radicals [*e.g.*, (12) and (13)] may play a significant part in the chemistry of free hydroxyl (compare Waters, *Trans. Faraday Soc.*, 1941, **37**, 779). Experimental support for this view is now forthcoming, since we have found that many substances, such as carboxylic acids and ketones, which are not themselves appreciably oxidised by free hydroxyl can noticeably depress the extent of oxidation of other substances (compare Fig. 5). This retarding effect can be accounted for by supposing that the added substance, in part, replaces the highly active hydroxyl radicals by other, more complex, radicals of very much lower oxidising power towards the organic substrate. The effect of this radical-transfer process may be evaluated as follows : If in the chain-oxidation

$$(12)\quad R\text{–}CO\text{–}OH + {\cdot}OH \rightleftharpoons R\text{–}CO\text{–}O{\cdot} + H\text{–}OH$$



$$(13)\quad \begin{matrix}R\diagdown \\ R'\diagup\end{matrix}C{=}O + {\cdot}OH \rightleftharpoons \begin{matrix}R\diagdown \\ R'\diagup\end{matrix}C\begin{matrix}\diagup OH \\ \diagdown O{\cdot}\end{matrix}$$

$$(14)\quad X\text{–}O\text{–}H + {\cdot}OH \underset{k_b}{\overset{k_a}{\rightleftharpoons}} X\text{–}O{\cdot} + H\text{–}OH$$

sequence of reactions (1)—(4) (p. S 15) we introduce also the equilibrium (14) and can write $[{\cdot}OX] = [{\cdot}OH] \,.\, K \,.\, [XOH]$, taking $[H_2O]$ as unity and presuming that both k_a and k_b are much greater than k_2 or k_3, then we have to consider also the consequent reactions

$$(2,b)\quad Fe^{++} + {\cdot}OX \xrightarrow{k_x} Fe^{+++} + (\colon\!OX)^-$$

$$(3,b)\quad XO{\cdot} + H\text{–}R \xrightarrow{k_y} XO\text{–}H + {\cdot}R$$

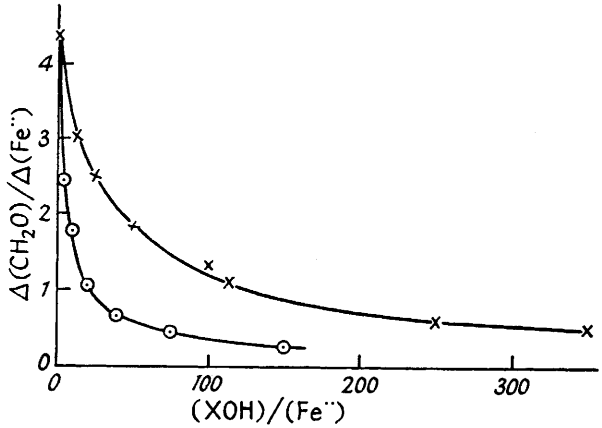
and thence, by evaluating [•OH], [•OX], and [•R] from instantaneous equilibria equations it can be derived that

$$\text{(H)}\quad \frac{-d[\text{H–R}]}{-d[\text{Fe}^{++}]} = \frac{[\text{H–R}]}{2[\text{Fe}^{++}]}\left(\frac{k_3 + k_yK[\text{XOH}]}{k_2 + k_xK[\text{XOH}]}\right)$$

$$\text{(J)}\quad \frac{-d[\text{H}_2\text{O}_2]}{-d[\text{H–R}]} = 1 + \frac{[\text{Fe}^{++}]}{[\text{H–R}]}\left(\frac{k_2 + k_xK[\text{XOH}]}{k_3 + k_yK[\text{XOH}]}\right)$$

which are the equivalent of (A) and (C) in which the ratio k_3/k_2 has been replaced by the more complex term in brackets. Analogous equations can be derived for the non-chain processes.

FIG. 5.



Retardation of the oxidation of formaldehyde.
× = *Effect of acetic acid.* ⊙ = *Effect of acetone.*

Now Fig. 6 for oxidations in which the ratio [R–H]/[Fe⁺⁺] has been kept constant indicates that the consumption ratio $\Delta[\text{H}_2\text{O}_2]/\Delta[\text{R–H}]$ for an alcohol or aldehyde varies with [XOH] in such a way that

$$\left(\frac{k_2 + k_xK[\text{XOH}]}{k_3 + k_yK[\text{XOH}]}\right)$$

is very closely of the form

$$\frac{(k_2 + \text{Const. [XOH]})}{k_3} \quad\text{or}\quad \frac{k_2}{k_3}\left(1 + \frac{C[\text{XOH}]}{k_2}\right)$$

This corresponds to the limiting case of $k_y = 0$, in which the secondary radical, •OX, is unable to attack the organic molecule R–H, though still is able to oxidise ferrous ions. This experimental fact would make it seem likely that the feebly active •OX radicals persist as such in water, and hence that equation (14) should be written not as an equilibrium but as a forward reaction with k_a comparable to k_3 in magnitude and $k_b = 0$. If this hypothesis is adopted, then from equations (1), (2,*a*), (2,*b*), (3), (4), and (14), we can derive the expression

$$\text{(L)}\quad \frac{-d[\text{Fe}^{++}]}{-d[\text{R–H}]} = \frac{2k_2[\text{Fe}^{++}]}{k_3[\text{R–H}]}\left(1 + \frac{k_a[\text{HOX}]}{k_2[\text{Fe}^{++}]}\right).$$

On the basis of equation (L) we can evaluate " transfer coefficients " k_a/k_2 which should be independent of the substrate which is being oxidised by the free hydroxyl still remaining in the solution. Table IV lists these values for the oxidation of *iso*propyl alcohol.

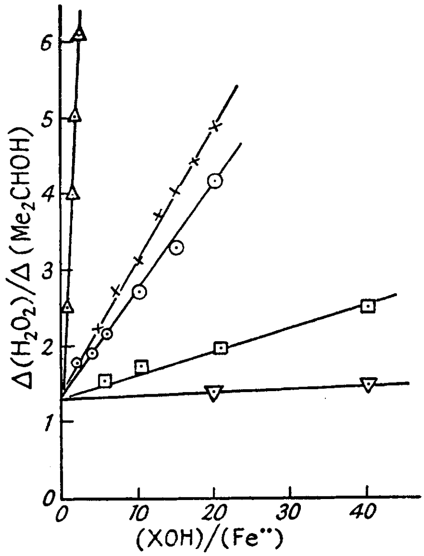
TABLE IV.

" *Transfer coefficients* ".

Acetic acid	0·0205	Acetone	0·162	Urea, sodium sulphate, potassium nitrate, ammonium sulphate	Nil
Malonic acid	0·0504	(Sodium) acetate	1·12		
Maleic acid	16·2	(Sodium) chloride	0·87		
Fumaric acid	33·3	(Potassium) bromide	21·6		

Evidently the large free radicals, such as the acetate radical, in which the " odd electron " can be spread over a wider region than is possible in free hydroxyl, are much less capable of attacking C–H bonds. Maleic and fumaric acids, which yield mesomeric radicals, $HO{-}CO{-}CH{=}CH{-}CO{-}O\cdot$, in which the olefinic system as well as both the hydroxyl groups may be concerned, exhibit this radical-transfer effect so markedly that the anticipated addition to the double bond could not be detected.

FIG. 6.



Retardation of the oxidation of iso*propyl alcohol.*

⊿ = *Maleic acid.* + = *Sodium acetate.*
⊙ = *Sodium chloride.* ⊡ = *Acetone.*
▽ = *Acetic acid.*

It is cogent in this connection to note that Kharasch and his colleagues (*J. Amer. Chem. Soc.*, 1943, **65**, 15; *J. Org. Chem.*, 1945, **10**, 386, 394) associate the enhanced oxidising powers of *hot* solutions of diacetyl peroxide, or lead tetra-acetate, with the breakdown of the feebly reactive acetate radical to the highly reactive free methyl radical, whilst similarly the reactivity of decomposing dibenzoyl peroxide is principally due to the free phenyl radical. These changes, it may be noted, are all decompositions of the $C{-}C{-}O\cdot$ structure which is also present in the unstable *tert.*-butyloxy-radical, $(CH_3)_3C{-}O\cdot$, of which mention has already been made.

The transfer coefficients of chloride and bromide anions are also of interest, for they accord with the relative orders of stability of the free halogen atoms. Evidently bromine atoms must have a fairly long free life in aqueous solution. Although the transfer coefficient noted for sodium acetate may in part denote the effect of changing the pH of the solution, it would be anticipated that acetate anions would lose electrons rather more easily than free acetic acid could lose a hydrogen atom.

Several oxidisable substances containing hydroxyl groups exhibit the effect of *self-retardation*, amongst which the resistance to oxidation by $\cdot OH$ of maleic and fumaric acids is an extreme example. For instance, crotonic acid shows the curious property of being oxidised more extensively in very dilute than in more concentrated solutions : $\Delta[R{-}H]/\Delta[Fe^{++}]$ first rises and then falls again as [R–H] is increased. At low concentrations hydroxylation of the double bond occurs since acetaldehyde can be detected as an oxidation product, whilst at high concentrations it may be suggested that the less active $CH_3{-}CH{=}CH{-}COO\cdot$ radical oxidises only Fe^{++}. Thiourea behaves in a similar way. Again, as we have previously noted (Merz and Waters, *loc. cit.*), the value of $\Delta[R{-}H]/\Delta[Fe^{++}]$ for alcohols approaches an upper limit at high alcohol concentrations. We had previously ascribed the latter effect to the increasing probability of chain termination by disproportionation, but an alternative view is that the " oxidising " properties which we have hitherto associated with the $\cdot OX$ radicals as well as the " reducing " properties which we have associated with the free carbon radicals are concerned.

If we take account of the mesomeric character of alcohol [or ether, compare (VI) and (VII)] radicals, and take the sequence of reactions (1), (2), (3), (4), and (2,*b*) (where OX = R) then we arrive at the more complicated equation (M), of which (A) is clearly one limiting form.

$$\text{(M)}\quad \frac{-\mathrm{d}[\mathrm{Fe}^{++}]}{-\mathrm{d}[\mathrm{R{-}H}]} = \frac{2k_2[\mathrm{Fe}^{++}]}{k_3[\mathrm{R{-}H}]}\left\{1 + \frac{(k_2k_x[\mathrm{Fe}^{++}] + k_3k_x[\mathrm{R{-}H}])}{(k_2k_4[\mathrm{H_2O_2}])}\right\}$$

This expression indicates (i) that at very high alcohol concentrations the consumption ratio $\Delta[\mathrm{Fe}^{++}]/\Delta[\mathrm{R{-}H}]$ should approach an upper limit, and (ii) that even at moderate alcohol concentrations the ratio $\Delta[\mathrm{R{-}H}]/\Delta[\mathrm{Fe}^{++}]$ should fall if the oxidiser (H_2O_2) is used in very dilute solution. Both these effects have, in fact, been noted by us already. In our previous paper we had ascribed the " dilution effect " (ii) to the participation of the equilibrium

$$\mathrm{R{-}CH_2{-}OH} + {\cdot}\mathrm{OH} \rightleftharpoons \mathrm{R{-}\dot{C}H{-}OH} + \mathrm{H{-}OH}$$

which is equivalent to suggesting that our mesomeric alcohol radical can oxidise water to hydroxyl. As our more recent discovery of radical transfer in aqueous solution has shown, the reversibility of this reaction is very unlikely. As yet, however, we have only examined the " self-retardation " and " dilution " effects in little more than a qualitative fashion, and can only make tentative surmises as to the causes of these minor deviations from the simple chain-reaction sequence which so clearly characterises the oxidations of alcohols by Fenton's reagent.

TABLE V.

Products of oxidations with the hydroxyl radical.

Substrate oxidised.	Products identified.	Derivative made or " spot test " used.
Methyl alcohol	CH_2O	D, H
Ethyl alcohol	$CH_3{\cdot}CHO$	D, H
n-Propyl alcohol............	$C_2H_5{\cdot}CHO$	H
*iso*Propyl alcohol	$CH_3{\cdot}CO{\cdot}CH_3$	H, Indigo
Me·CH(OH)·CH₂·CH₂·OH	$CH_3{\cdot}CO{\cdot}CH_2{\cdot}CO_2H$, $CH_3{\cdot}CH{:}CH{\cdot}CHO$	(H colour), Fe^{3+} colour, H
Diethyl ether	$CH_3{\cdot}CHO$	D
Dioxan........................	CH_2O, $H{\cdot}CO_2H$, oxalic acid	D, Resorcinol test; Ca salts
Tetrahydrofuran............	{CH_2O (from both) }	D, H
Tetrahydropyran	{Acids $HO_2C{\cdot}[CH_2]_n{\cdot}CHO$}	H colours
Glycollic acid	(Not established)	
Lactic acid	$CH_3{\cdot}CO{\cdot}CO_2H$	H
Thioglycollic acid	(Not established)	
Glycine........................	CH_2O	D
Alanine........................	$CH_3{\cdot}CHO$	Nitroprusside test
tert.-Butyl alcohol	$CH_3{\cdot}CO{\cdot}CH_3 + CH_2O$	D, H, Indigo
tert.-Amyl alcohol	{$CH_3{\cdot}CO{\cdot}CH_3 + CH_3{\cdot}CHO$} {$C_2H_5{\cdot}CO{\cdot}CH_3 + CH_2O$}	D, H, Nitroprusside, piperidine, and resorcinol tests
$C_6H_5{\cdot}CMe_2{\cdot}OH$	$C_6H_5{\cdot}CO{\cdot}CH_3 + CH_2O$ (*no* phenol)	D, H
Chloral hydrate	(Not established)	
Pinacol........................	$CH_3{\cdot}CO{\cdot}CH_3$	H, Indigo
Ethylene glycol	$CH_2(OH){\cdot}CHO + CH_2O$ (about 1%)	H (osazone), D (weighed)
2 : 3-Butylene glycol	$CH_3{\cdot}CO{\cdot}CO{\cdot}CH_3$	Ni glyoxime
1 : 2-Butylene glycol	$C_2H_5{\cdot}CO{\cdot}CO_2H$	H (red soln. in alkali), Fe^{3+} colour
Methyl acetate	CH_2O	D
Ethyl acetate	$CH_3{\cdot}CHO$	Piperidine and nitroprusside tests
*iso*Propyl acetate	$CH_3{\cdot}CO{\cdot}CH_3$	Indigo
Propionic acid...............	$CH_3{\cdot}CHO$	D, Piperidine and nitroprusside tests
n-Butyric acid	(Not established)	
*iso*Valeric acid...............	$Me_2CH{\cdot}CHO$	H
Succinic acid	(Not established)	
Adipic acid	(Not established)	
Crotonic acid	$CH_3{\cdot}CHO$	D, Piperidine and nitroprusside tests
Diethylamine	$CH_3{\cdot}CHO$	„ „
Triethylamine...............	$CH_3{\cdot}CHO$	„ „
Pyridine	(Blue solution)	

Notes.—D = Dimedone derivative; H = 2 : 4-dinitrophenylhydrazone.

EXPERIMENTAL.

The oxidations were carried out at room temperature as follows: 10 ml. of 0·1N-ferrous ammonium sulphate in 0·5N-sulphuric acid were mixed with the appropriate volume of a standard aqueous solution of the organic substrate and diluted with distilled water to a total of 45 ml. 5·0 Ml. of 0·1N-hydrogen peroxide were then added with shaking, and after ½ hour or longer the [Fe^{+++}] content of the product was determined colorimetrically by means of either a Duboscq colorimeter or (in later experiments) a "Spekker" photo-absorptiometer. For this determination the reaction mixture was diluted to 250 ml.; a 5-ml. portion was then mixed with 5 ml. of 10% potassium thiocyanate and diluted to 100 ml. with 0·2N-sulphuric acid, a similarly prepared mixture to which no organic substrate had been added serving as the reference blank.

In radical-transfer experiments the oxidisable organic substrate was invariably added as 10 ml. of a 0·2M-solution. Formaldehyde was used for the data of Fig. 5, and *iso*propyl alcohol for those of Table IV and Fig. 6.

For the data plotted in Fig. 3, the Hg^+ formed was precipitated as mercurous chloride with hydrochloric acid, filtered off, and then determined iodometrically. The Fe^{+++} was determined, as usual, in the filtrate. $\Delta[Me_2CH{\cdot}OH]$ was calculated from the relationship

$$\Delta[Me_2CH{\cdot}OH] = \Delta[H_2O_2] - \Delta[Fe^{++}] + \Delta[Hg^{++}].$$

Table V summarises the identifications of organic oxidation products. Most of these were separated by fractional distillation. Solid derivatives were in every case crystallised until their mixed m. p.s with authentic specimens showed no depression.

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