rate constant at the last point significantly lower than at the first point.

Mechanism.—Previous studies of base-catalyzed hydrolysis of esters and amides,[9] ammonolysis of esters[2,10] and ester exchange[11] have indicated that in the case of acyl–oxygen fission all can probably be included in one mechanistic scheme. The O^{18} studies of Bender (see ref. 10a) show that this must involve, for ester hydrolysis at least, an addition of base to the carbonyl group giving an ionic intermediate. For the hydrolysis of amides the O^{18} exchange evidence is permissive and the lack of salt effect[10d] is in accord with that interpretation. The data recorded here permit the inclusion of amide alcoholysis in the same group.

In support of this conclusion the following evidence may be cited. The rate is first order in anilide and first order in base and more rapid when the more basic[12] catalyst ethoxide ion replaces methoxide. Jaffé[13] using the data accumulated by Verkade and his co-workers[14] has shown that a Hammett plot of the rate *versus* σ-constants for various substituents is linear with $\rho = +1.723$. This ρ-value is in good agreement with that for hydroxyl ion catalyzed hydrolysis of ethyl benzoates (+2.37)[15] and that for methoxide ion catalyzed ester exchange of menthyl benzoates (+2.530).[11] Also using the data of Verkade[14] and assuming a temperature of 65° for his work, it is possible to approximate (Table III) energies and entropies of activation for this reaction. The results agree reasonably with values for other reactions in the series as is illustrated by Table IV.

TABLE III

ESTIMATED ENERGIES AND ENTROPIES OF ACTIVATION FOR METHANOLYSIS OF *o*- AND *p*-NITROACETANILIDES

	o-NO₂	*p*-NO₂
k_2(30°)	1.10×10^{-2}	1.34×10^{-4}
k_2(65°)[a]	19×10^{-2}	80×10^{-4}
k_a, kcal./mole	16.7	23.8
$\Delta S^{\pm}$(303°), cal./deg.-mole	−12.4	2.3

[a] Data from Verkade, *et al.*[14]

TABLE IV

ACTIVATION ENERGIES AND ENTROPIES

Reaction	E_A, kcal./mole	$\Delta S^{\pm}$, cal./deg.	Refer.
p-Nitrobenzamide and OH^-	16.1	−30.0	9c
Menthyl *o*-nitrobenzoate and MeO^-	16.9	−20.4	11
Menthyl *p*-nitrobenzoate and MeO^-	14.8	−19.9	11
Ethyl *o*-nitrobenzoate and OH^-	14.5	−20.4	15
Ethyl *p*-nitrobenzoate and OH^-	14.8	−14.1	15
p-Nitrophenyl acetate and OH^-	11.0	−17.4	15

The spectral results make it seem probable that the alcoholysis reaction is preceded by a rapidly established equilibrium between the catalyst and the substrate. The anilide ion produced thus

$$ArNHCOCH_3 + OR^- \rightleftarrows ArNCOCH_3^- + ROH$$

would be expected to show a more intense absorption than the anilide itself. That the reaction is reversible was illustrated by reappearance of the original spectra after neutralization of the base. If the equilibrium constant for the reaction is small, it can be shown easily that the equilibrium will not influence the nature of the kinetic results, which seems to be the case under the conditions of our study.

CORVALLIS, OREGON

---

[9] For discussion and references see (a) J. Hine, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1956, pp. 266–268, 295–298; also (b) A. Bruylants, *et al.*, *Bull. soc. chim. Belges*, **60**, 191 (1951); **63**, 140 (1954); (c) K. Laidler and I. Meloche, THIS JOURNAL, **73**, 1712 (1951); (d) J. Packer, A. Thompson and J. Vaughn, *J. Chem. Soc.*, 2601 (1955); (e) C. Bunton and O. Spatcher, *ibid.*, 1079 (1956).

[10] F. H. Wetzel, J. G. Miller and A. R. Day, THIS JOURNAL, **75**, 1150 (1953), and earlier papers.

[11] R. Taft, Jr., M. Newman and F. H. Verhoek, *ibid.*, **72**, 4511 (1950).

[12] J. Hine and M. Hine, *ibid.*, **74**, 5266 (1952).

[13] H. Jaffé, *Chem. Revs.*, **53**, 191 (1953).

[14] P. E. Verkade, *et al.*, *Rev. trav. chim.*, **71**, 545, 1245 (1952); **70**, 127 (1951); **69**, 1393 (1950); **68**, 88 (1949); **67**, 411 (1948).

[15] E. Tommila and C. N. Hinshelwood, *J. Chem. Soc.*, 1801 (1938).

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY AND RADIATION LABORATORY, UNIVERSITY OF CALIFORNIA, BERKELEY]

Chemistry and Spin Resonance Spectroscopy of Radicals from Thioaromatic Compounds[1]

BY ANTONINO FAVA,[2] P. B. SOGO AND MELVIN CALVIN

RECEIVED OCTOBER 4, 1956

The deeply colored solutions which result when a wide variety of aromatic sulfides are dissolved in concentrated H_2SO_4 are shown to contain free radicals having the thianthrene nucleus. These result from the condensing and oxidizing properties of concentrated H_2SO_4. A simplified system in trifluoroacetic acid has been studied in detail and the radical shown to have an oxidation level lying between the monoxide and dioxide of thianthrene. An interpretation of the spin resonance spectrum permits a choice of specific radical structure.

Recently it has been reported by Hirshon, Gardner and Fraenkel[3] that the sulfuric acid solutions of certain sulfur-containing aromatic compounds show electron spin resonance and contain free radicals in substantial amounts. The substances studied include phenyl and *p*-tolyl disulfides, the corresponding thiols and diphenylene sulfide (Thianthrene). No definite hypothesis was made concerning the nature of the free radicals, although the authors seemed to suggest the formation of a biradical.

---

[1] The work described in this paper was sponsored in part by the U. S. Atomic Energy Commission and in part by the Chemistry Department, University of California, Berkeley, California.

[2] On leave from the University of Padova, Italy. International Cooperation Administration Fellow, 1954–1956.

[3] J. M. Hirshon, D. M. Gardner and G. K. Fraenkel, THIS JOURNAL, **75**, 4115 (1953).

More recently Wertz and Vivo[4] have investigated these compounds also from the standpoint of the spin resonance. These authors found that if the solutions of the disulfides and thiols are made and examined in the absence of air, two different radicals may be observed. One of these is quite stable and appears to be identical with that obtained from thianthrene.

The present investigation was carried out with the object of identifying the stable free radical and determining the nature of the reaction between aromatic disulfides and sulfuric acid.

Experimental

Eastman Kodak Co. thianthrene recrystallized from acetic acid, m.p. 158°, was used for the preparation of the oxygenated derivatives. Thianthrene monosulfoxide was prepared by oxidation of thianthrene in acetic acid[5] with dilute HNO_3 and was recrystallized from alcohol; m.p. 143°. The isomeric disulfoxides were prepared according to the method of Fried and Vogt[6]; *trans*, m.p. 284°; *cis*, m.p. 249°.

Trifluoroacetic acid and 98% reagent grade sulfuric acid were used. Optical spectra were observed with a Cary model 14 spectrophotometer. The electron paramagnetic resonance spectra were taken at a frequency of 9.25 K. mc./s. on a recording differentiating spectrometer (similar to one developed by Beringer and Castle[7]), using a transmission cavity and bolometer detection.

Results and Discussion

The solutions of thianthrene in concentrated H_2SO_4 are characterized by two absorption bands in the near infrared with maxima at 10,500 and 9200 Å., one band in the visible at 5460 Å. and three bands in the ultraviolet (Fig. 1).

In the range of concentration between 10^{-3} to 10^{-4} mole/liter, Beer's law is not obeyed in the visible and near infrared region. A tenfold dilution of the concentration produced only an eightfold decrease in absorption. At concentrations around 5×10^{-5}, Beer's law is obeyed.

In the range of concentrations in which Beer's law is not followed, the optical densities at the maxima in the near infrared and visible have the same variation with concentration as the spin resonance signal. It was recognized, therefore, that these bands are characteristic of the radical.

Although in solutions of diphenyl disulfide in H_2SO_4 (prepared in air) a new weak band is detectable at 7250 Å., absorption of light in the near infrared and visible wave length region characteristic of the radical is present. The amount of radical as determined by paramagnetic resonance was found to be proportional to the optical density at 5460 Å.

These findings confirmed the observation that the radical obtained from thianthrene and that obtained from diphenyl disulfide in air are identical. All subsequent experiments, therefore, were carried out on thianthrene and its derivatives.

It was interesting to see if thianthrene, after solution in H_2SO_4, could be recovered as such by diluting the acid. A 2.5×10^{-3} *M* solution of thianthrene prepared by shaking three days at room temperature was poured on crushed ice and the suspension exhaustively extracted with cyclohexane. The ultraviolet spectrum of the cyclohexane layer was compared with the spectrum of thianthrene in the same solvent. The two spectra coincided but the amount of thianthrene was approximately 50% of the total. The aqueous layer showed a strong absorption band at 2600 Å.

This experiment showed that dissolving in H_2SO_4 transforms some of the thianthrene into a water (aqueous acid) soluble product, probably one of the oxides of thianthrene, which are known to be slightly soluble in H_2O and possibly some sulfonation products. That an oxidation is involved is supported by the fact that a small amount of SO_2 is evolved during solution.

After this experiment it was interesting to see if the radical would be formed by a solution of thianthrene in other acids. Pure trifluoroacetic acid was found capable of dissolving thianthrene with the formation of the radical. The optical spectrum in CF_3–COOH showed the characteristic bands in the near infrared and visible (ultraviolet cannot be observed below 3000 Å.). The spectrum, which is much better resolved than in H_2SO_4, is shown in Fig. 1. The band at 10,500 Å. is a doublet with maxima at 10,600 and 10,250 and an additional weak band was observed at 8000. In the same range of concentrations studied in H_2SO_4, Beer's law fails in this system as well.

It appeared from the first quantitative experiments that the amount of the radical formed in trifluoroacetic acid depended on the amount of oxygen available, although even when the solution was made up *in vacuo* and with degassed reactants, a certain amount of the radical (about 5% of the total thianthrene) always was formed. The heating, in a sealed vessel (80°) of the solution prepared *in vacuo*, causes first a small increase in the concentration of the radical and subsequently a gradual decrease. At equilibrium less than 2% of the initial amount of radical (about $^1/_{100}$ of the total thianthrene) is left. These experiments suggested that oxygen and oxygenated derivatives of thianthrene play an essential role in the formation of the radical. Since H_2SO_4 is simultaneously an acid and an oxidizing agent, the difficulty of the interpretation of the experiments carried out in H_2SO_4 was recognized. To separate the two effects, subsequent investigations were extended to the oxides of thianthrene *in vacuo* using trifluoroacetic acid as a solvent.

Six oxides of thianthrene are described in the literature: monosulfoxide, *cis*- and *trans*-disulfoxides, monosulfone, monosulfoxide monosulfone and disulfone. Except for the disulfone, all are known to give colored solutions in sulfuric acid.

We have not been able to prepare, according to the directions of the workers who prepared them originally,[5a] the two monosulfones. We have not been able to find in the literature indications that these compounds have been prepared by others: rather it has been found that other investigators[8] failed in preparing them.

The solutions in trifluoroacetic acid (in the absence of air) of the monosulfoxide and the two disul-

---

[4] J. E. Wertz and J. Vivo, *J. Chem. Phys.*, **23**, 2193 (1955).

[5] (a) K. Fried and W. Vogt, *Ann.*, **381**, 312 (1911); (b) T. P. Hilditch, *J. Chem. Soc.*, **97**, 2579 (1910).

[6] K. Fries and W. Vogt, *Ber.*, **44**, 756 (1911).

[7] R. Beringer and J. G. Castle, *Phys. Rev.*, **78**, 581 (1950).

[8] J. Boeseken and A. T. H. van der Meulen, *Rec. trav. chim.*, **55**, 929 (1936).

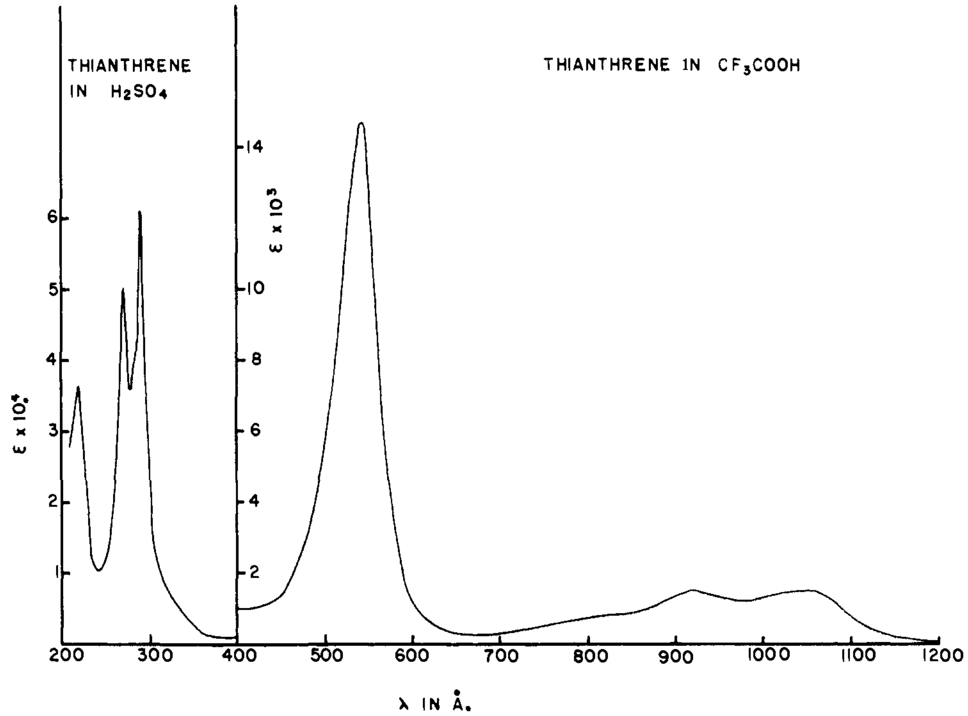


Fig. 1.—Absorption spectrum of thianthrene radicals in acid solution (10^{-5} M in H_2SO_4; 10^{-4} M in CF_3CO_2H).

foxides are colorless and no radical is present. However, very slowly at room temperature, and more rapidly at 80°, the usual pink color appears and the solutions show paramagnetic resonance. At 80° an equilibrium value is obtained in about 15–20 hours. The formation of the radical is much faster for the monosulfoxide than for the two disulfoxides. Also the amount formed at equilibrium is higher. For 1.85×10^{-3} M solutions, 15% of the monosulfoxide and about 3% of either of the disulfoxides is transferred into the free radical.

These results suggested that the free radical is an intermediate state of oxidation between two oxides, possibly mono- and disulfoxide. This hypothesis was tested by measurements of the relative intensities of the spin resonance absorption obtained at equilibrium from different proportions of mono- and disulfoxide mixed in solution. The results are given in Table I.

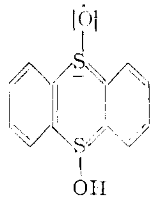
TABLE I

CONCENTRATIONS IN MOLE/LITER OF THIANTHRENE (T), MONOSULFOXIDE (M), DISULFOXIDE (D) AND FREE RADICAL (R)

$(T) \times 10^3$	1.85	..	..	1.85	..	..	..
$(M) \times 10^3$	..	1.85	..	..	1.85	1.85	1.85
$(D) \times 10^3$	..	..	1.85	1.85	1.85	3.70	7.40
$(R) \times 10^3$	0.002	0.28	0.06	0.59	0.80	1.33	1.75
$K = (R\cdot)^2/(M)(D)$					.31	0.49	0.48

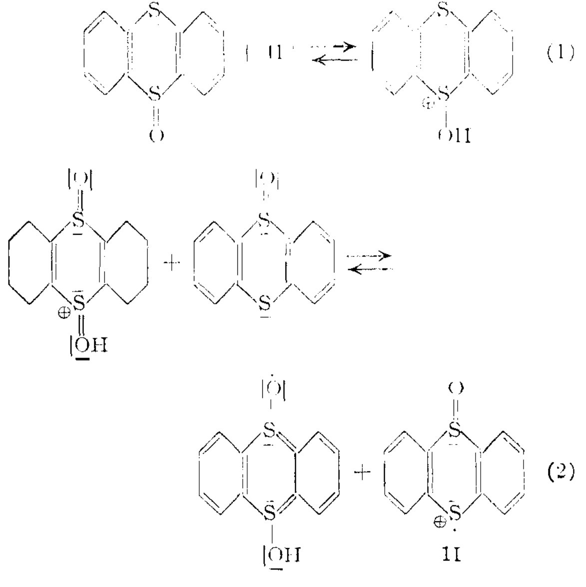
These results support the view that mono- and disulfoxide are in equilibrium with one another and with the radical, the latter being an intermediate state of oxidation between the two.

The hypothesis is made that this radical is of the form

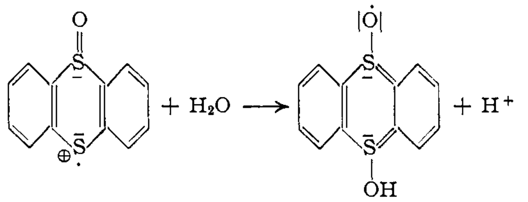


I

and that it is formed through the steps

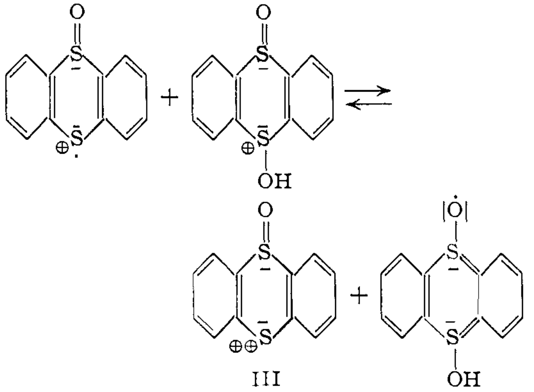


II may now react with a water molecule according to



(3)

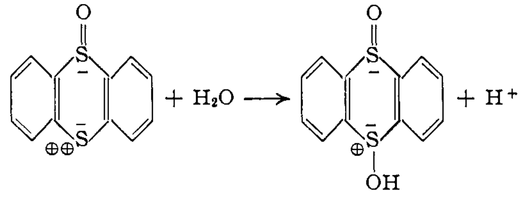
Alternately II may reduce another disulfoxide molecule acccording to



III

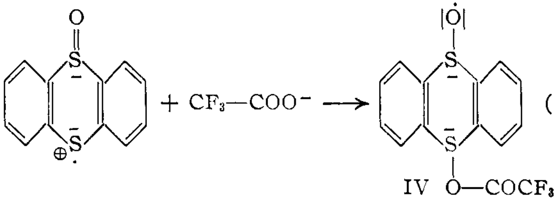
(4)

followed by



(5)

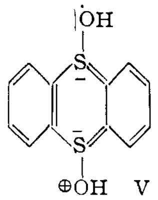
It has been assumed so far that there is always enough water to perform reaction 3 (or 5). However, even in an anhydrous acid similar reactions 6 may be devised



IV

(6)

and in sulfuric acid there could be a corresponding derivative. There exists also the possibility that the radical has a symmetric structure such as



V

or the equivalent with the acyl groups instead of the protons. The relative merit of these structures will be discussed later on the basis of the spin resonance spectrum.

Whatever is the actual structure of the radical, an equilibrium is established such that

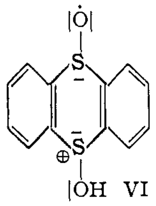
$$\frac{(R\cdot)^2}{(M) \times (D)} = K \qquad (7)$$

If in the first approximation[9] it is assumed that

$$(M) = (M)_0 - {}^1/_2(R\cdot) \qquad (8)$$
$$(D) = (D)_0 - {}^1/_2(R\cdot)$$

the equilibrium constant may be calculated. This has been done for those experiments of Table I involving both monosulfoxide and disulfoxide and the values obtained are shown there.

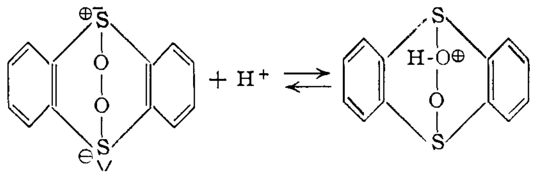
It may be mentioned that a biradical structure, of the type suggested by Hirshon, Gardner and Fraenkel[3] for the bianthrone



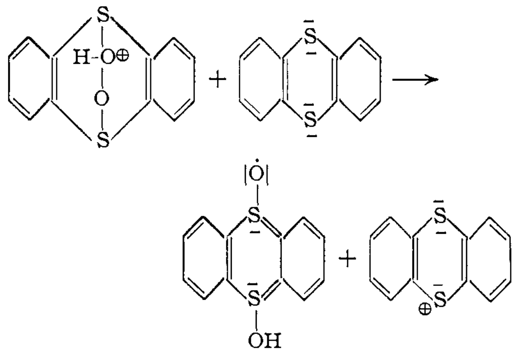
VI

can be ruled out here; in fact VI could be most easily formed from disulfoxide plus H$^+$, whereas it has been shown that both disulfoxides hardly form any of the radical in trifluoroacetic acid.

It is of interest to consider the fact that solutions of thianthrene in trifluoroacetic acid even when made up *in vacuo* exhibit a considerable amount of the radical as soon as they are made up. It was thought at first that an improper degassing of the reagent was the cause, but repeated careful experiments showed that no matter how the reagents as such are degassed, about 5% of the total thianthrene is transformed immediately into the free radical upon going into solution. We were, therefore, forced to admit that a certain amount of oxygen is so strongly absorbed onto the thianthrene that it cannot be removed by simply decreasing the pressure. Probably, a small fraction of thianthrene is in the form of a transannular peroxide. The fact that the heating of these types of solutions causes the disappearance of most of the radical (exactly the opposite of what happens when monosulfoxide and disulfoxide are initially present) suggests that here the radical is formed first; subsequently it disproportionates into mono- and disulfoxide until the equilibrium is attained. A tentative explanation for this initial formation of the radical from the transannular peroxide could be



(9)

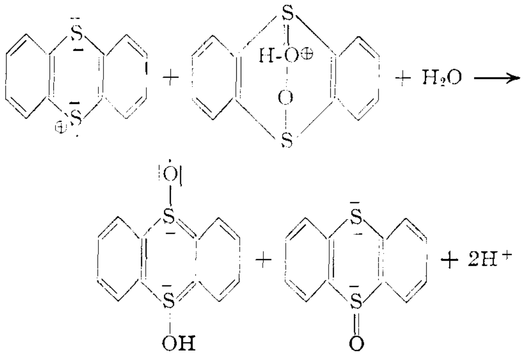


(10)

---

[9] In this approximation the amounts of mono- and disulfoxide that might have been transformed through other reaction, disproportionation and/or formation of radical dimers, are neglected.

VII could, in the following reaction, reduce another molecule of transannular peroxide, being itself transformed into monosulfoxide by the intervention of a water molecule. It seems quite reasonable that the above sequence would take place very rapidly since the transannular peroxide is certainly in a much higher energy state than the isomeric disulfoxide.


(11)

The Spin Resonance Spectrum.—As already observed by the previous workers, the free radical spin resonance spectrum of this series of compounds consists of five lines with approximate intensity ratios 1:4:6:4:1. A typical derivative curve is presented in Fig. 2. The *g*-value for the center of the pattern was found to be 2.009 ± 0.002 by a comparison with the resonance from crystalline DPPH for which we assumed the *g*-value of 2.0036. The spacing of the five component lines was 1.5 ± 0.1 gauss.

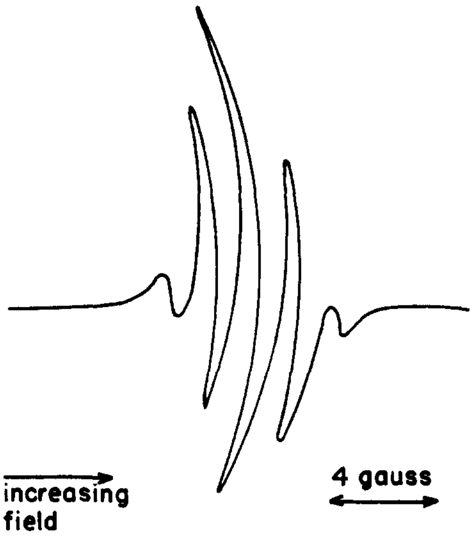
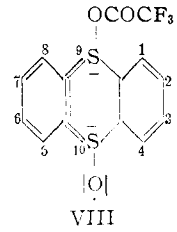


Fig. 2.—Typical derivative curve of the free radical spin resonance (1.75×10^{-3} mole/liter of radical in CF_3CO_2H).

Assuming that the fine structure is due to the interaction of the magnetic moment of the unpaired electron with nuclear magnetic moments, it is deduced that the free electron interacts with only four protons. Since at least eight ring protons are present in each of the radicals proposed above (I, IV, V), at least nine lines should be expected, if all of the protons were equivalent.

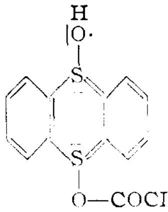
However, if the structure radical is an asymmetric one, such as IV, in which only the oxygen on position 10 is bound to an acid residue, it is easily seen, by writing down all the resonance formulas not involving separation of charge, that the free electron may be found on 6 of the 12 carbon atoms, four of which (2, 4, 5 and 7) are bound to hydrogen atoms. This situation is similar to the non-equivalence of the ring positions observed by Tuttle, Ward and Weissman[10] in naphthalene negative ion. If the acid residue is bound to the oxygen on position 9 (VIII), the free electron would interact primarily only with protons on 1, 3, 6 and 8.


VIII

Since IV and VII are tautomers the free electron interacts significantly with only four protons at a time.

If this interpretation of the spin resonance spectra is correct, an asymmetric radical such as IV with no additional protons would be the only one present in large amount.

A search for free radicals of low abundance was made. Three groups of lines with *g*-values sufficiently different from 2.009 were observed. The heights of these lines were about $^1/_{500}$ of the height obtained for the principal radical above, while the widths were found to be about the same. Two groups consisting of five lines each were found to be symmetrically displaced on either side of the principal resonance by a field of 16 ± 1 gauss. The intensity ratio, hyperfine splitting and general asymmetric appearance of the individual lines in the two groups are the same as for the resonance of the principal radical. The intensity of the low field resonance is about twice that of the high field resonance. It is suggested that these two groups of lines originate from the radical formed by the addition of a proton to VIII (or IV). The 32 gauss splitting between the groups is due to hyperfine interaction between the proton on the oxygen and the electron. This may be compared with the 24 gauss splitting which has been attributed to the methyl proton in the radical dimesitylmethyl.[11] One would expect, because of the greater electronegativity of the oxygen, that the unpaired electron would spend a larger proportion of its time


IX

---

[10] T. R. Tuttle, Jr., R. L. Ward and S. I. Weissman, *J. Chem. Phys.*, **25**, 189 (1956).

[11] H. S. Jarrett and G. J. Sloan, *ibid.*, **22**, 1783 (1954).

there as compared to the ring positions. In this interpretation, one should expect equal intensities for the two groups, which was not observed. However, a *g*-value difference of 1% for free radicals with otherwise identical spectra seems more unlikely than an explanation such as above. In this way a *g*-value for the center of the two groups falls at $g = 2.009 \pm 0.002$.

The third group of lines has a *g*-value only about 0.025% higher than the principal resonance with the result that almost all of it is obscured. Three lines are observed on the low field side and two lines appear on the high field side of the principal resonance. The splitting is again about 1.5 gauss and the total spread about 15 gauss.

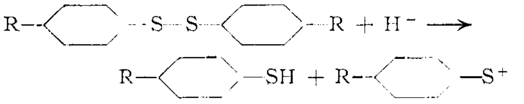
The intensities of these lines are too small to allow their inclusion as part of the main five-line spectra so we have ascribed them to another free radical with an undetermined number of lines. The superposition of the relatively very intense spectra upon most of this resonance together with poor signal-to-noise ratio prevented a detailed study of this radical. However, a division of the 15 gauss spread by the 1.5 gauss splitting suggests the presence of the nine lines which would be expected of a radical of type V with fluoroacetate replacing the hydroxylic groups.

The Reaction between Aromatic Disulfides and Sulfuric Acid.—It was interesting to investigate whether other aromatic disulfides besides those reported by previous workers[3,4] dissolve in concentrated sulfuric acid with the formation of free radicals. A number of substituted aromatic disulfides have been investigated and the results are

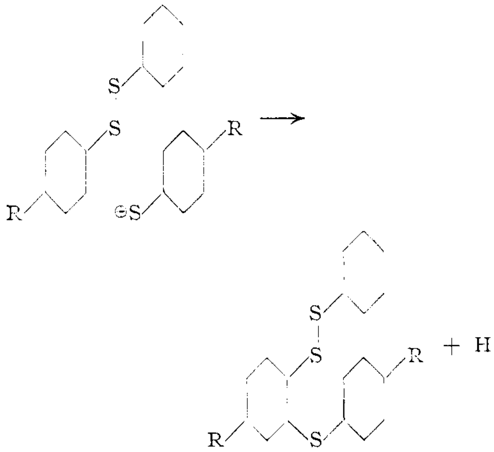
Disulfide	Color of the H_2SO_4 soln.	Approx. abundance of free radical, %
Diphenyl	Purple	5–10
p-Tolyl	Purple	~10
o-Tolyl	Blue-green	~4
p-Nitro	Red	. .
o-Nitro	Yellow	. .
p-Dimethylamino	Colorless	. .
o-Carboxy	Orange-red	1
p-Chloro	Does not dissolve	

In analogy with what has been established for diphenyl disulfide it may be assumed that whenever free radicals are detected, a derivative of thianthrene has been formed from the corresponding disulfide. An analysis of the above results shows that with the exception of the *o*-carboxy derivative, the disulfides capable of reacting with sulfuric acid to give thianthrenes are those which contain electron-releasing substituents in the ring. When electron-attracting substituents are present, the formation of thianthrenes is prevented.

The mechanism shown, consistent with the above observation, is proposed

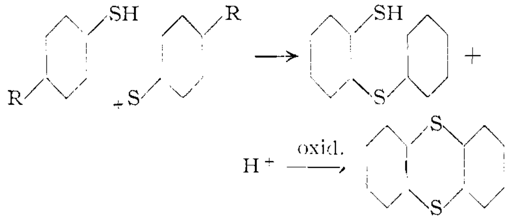


(12)



(13)

with later ring closure and formation of the corresponding thiophenol. The electrophilic attack may occur not only on the disulfide but also on the thiophenol. In this case ring closure occurs with simultaneous oxidation performed by the sulfuric acid.



(14)

In this mechanism it is seen that the formation of thianthrenes is determined by the rate of reaction 12, 13 or 14. Naturally it is impossible to state which step is the rate-determining one, but whichever it may be, it is seen easily that all reactions are favored by electron-releasing and hindered by electron-withdrawing substituents. The only exception to this behavior is offered by the carboxy derivative and this may be accounted for reasonably by assuming that in sulfuric acid a partial decarboxylation occurs, making possible the subsequent electrophilic attack (13).

BERKELEY 4, CALIFORNIA