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Mechanism of Oxygen Exchange Reaction of Diaryl and Alkyl Aryl Sulfoxides in Sulfuric Acid*

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The mechanistic investigations of the concurrent oxygen exchange and racemization reactions of ^{18}O-labeled and optically active methyl *p*-tolyl sulfoxide (I) and phenyl *p*-tolyl sulfoxide (II) in sulfuric acid of various concentrations have been carried out by the determination of detailed kinetics. The kinetic data were analyzed in the lights of k_{ex}/k_{rac} values, activation parameters, correlations between the rates and acidity functions, solvent isotope effects, polar effects of substituents, etc.. A gradual change of mechanism of the reaction from an S_N2 type process to a predominant A-1 like one (in above 95% sulfuric acid) with the increase of the concentration of sulfuric acid was observed. All these observations suggest that the A-1 like reaction takes place through a cation radical $(\text{—}\overset{+\cdot}{S}\text{—})$ or a dication $(\text{—}\overset{++}{S}\text{—})$ intermediate at higher concentrations of sulfuric acid, while the S_N2 type reaction that involves a water molecule as nucleophile in the rate-limiting step is the predominant path in less concentrated sulfuric acid.

The concurrent oxygen exchange and racemization reactions of optically active and ^{18}O-labeled (*on sulfoxide oxygen*) sulfoxides are advantageous to study the nature of nucleophilic substitution on the trivalent sulfur atom, which is one of the important subject in organic sulfur chemistry, and have attracted considerable attentions in recent years. Several studies have been reported on the acid-catalyzed oxygen exchange and racemization reactions of sulfoxides.[1–7] Among these, the concurrent oxygen exchange and racemization reactions of diaryl sulfoxides in concentrated sulfuric acid is the most extensitvely studied, since it was the first example discovered.[8–10] When ^{18}O-labeled and optically active sulfoxides are dissolved in a concentrated sulfuric acid, one observes the concurrent oxygen exchange and racemization of the sulfoxide.

However, in order to clarify the over-all feature of these reactions, in which the mode of the reaction displays a gradual change with the change of the concentration of sulfuric acid, more detailed kinetic investigations are desirable especially since it has recently been shown that the sulfoxides do not behave as Hammett bases.[19,20,31]

We now have investigated the oxygen exchange and racemization reactions of methyl *p*-tolyl sulfoxide (I) as well as phenyl *p*-tolyl sulfoxide (II) much more in detail in sulfuric acids of various concentrations. And a reinvestigation of our previous results[9,10] has also been carried out.

Results and Discussion

Careful kinetic experiments have been conducted mainly on the rates of concurrent oxygen exchange and racemization reactions of methyl *p*-tolyl sulfoxide (I), 1.08 *atom*%-^{18}O $[\alpha]^{23}_{D}+146°$ (Acetone), and phenyl *p*-tolyl sulfoxide (II), 1.12 *atom*%-^{18}O $[\alpha]^{20}_{D}+27.6°$ (EtOH), in sulfuric acid of various concentrations ranging from 52.4% to 94.8% at several temperatures. These results are listed in Table 1.

When the reaction proceeds through the A–1 type

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* Oxygen Exchange Reaction of Sulfoxides in Sulfuric Acid. Part VII (Sulfoxides 47).

1) a) K. Mislow, T. Simmons, J. T. Mellilo, and A. L. Ternay, Jr., *J. Amer. Chem. Soc.*, **86**, 1452 (1964). b) K. Mislow, *Record of Chemical Progress*, **28**, 217 (1967).

2) H. Yoshida, T. Numata, and S. Oae, This Bulletin, **44**, 2875 (1971).

3) G. Modena, U. Quintily, and G. Scorrano, *J. Amer. Chem. Soc.*, **94**, 202 (1972).

4) S. Oae, M. Yokoyama, and M. Kise, This Bulletin, **41**, 1221 (1968).

5) N. Kunieda and S. Oae, *ibid.*, **41**, 1025 (1968).

6) a) S. Oae, N. Kunieda, and W. Tagaki, *Chem. & Ind.*, **1965**, 1790. b) N. Kunieda, K. Sakai, and S. Oae, This Bulletin, **42**, 1090 (1969).

7) M. Kise and S. Oae, *ibid.*, **43**, 1804, 1416, 1426 (1970).

8) S. Oae, T. Kitao, Y. Kitaoka, and S. Kawamura, *ibid.*, **38**, 546 (1965).

9) S. Oae and N. Kunieda, *ibid.*, **41**, 696 (1968).

10) N. Kunieda and S. Oae, *ibid.*, **42**, 1324 (1969).

TABLE 1. RATE CONSTANTS FOR THE OXYGEN EXCHANGE AND RACEMIZATION REACTIONS OF p-CH_3-C_6H_4-SO-R IN SULFURIC ACID

R;	H_2SO_4 (wt%)	Temp. (°C)	10^4k_{rac} (s^{-1})	10^4k_{ex} (s^{-1})	10^4k_1 (s^{-1})[e]	Method[b]
CH_3[a]	94.8[c]	5	11.2 ±0.5	9.63 ±0.12		A
	94.2	12.5	24.5 ±0.1		24.5	B
	94.2	10	17.8 ±0.2		17.8	B
	94.2	5	9.61 ±0.21		9.61	B
	93.8	20	38.9 ±1.2		38.9	B
	93.8	15	21.5 ±0.3		21.5	B
	93.8	10	11.8 ±0.2		11.8	B
	93.8	5	6.37 ±0.23		6.37	B
	93.1	12.5	17.7 ±0.07		17.7	B
	92.7[c]	30	81.7 ±1.5	57.2 ±2.2		A
	91.9	12.5	14.0 ±0.8		14.0	B
	89.8	12.5	7.72 ±0.73		7.72	B
	89.5	30	40.3 ±1.5		40.3	B
	83.5	30	8.50 ±0.10		8.52	B
	80.2	30	3.55 ±0.11		3.56	B
	74.6	30	1.19 ±0.06		1.20	B
	68.8	30	0.440±0.010		0.456	B
	64.7	30	0.266±0.002		0.285	B
	64.7	40	0.703±0.015		0.753	B
	64.7	47	1.20 ±0.02		1.29	B
	58.5	30	0.145±0.05		0.170	B
	58.5[c]	30	0.263±0.022	0.126±0.010		A
	55.0	30	0.105±0.003		0.135	B
	52.4	30	0.074±0.001		0.104	B
C_6H_5[d]	95.9	5	4.20	4.07	4.20	A
	95.5	5	3.32	3.22	3.32	A
	91.0	5	0.891	0.705	0.891	A
	86.9	30	2.91	1.89	2.91	A
	83.4	30	1.39	0.900	1.40	A
	80.5	30	0.880	0.458	0.884	A
	75.4	60	6.95	3.40	7.06	A

a) Sulfoxide=0.03 mol/l. b) Method of the rate measurement; see Experimental section. c) Sulfoxide=0.1 mol/l. d) Data of our previous experiment.[10] e) The pK_{SH^+} values used here are, for phenyl p-tolyl sulfoxide (−2.39), for methyl p-tolyl sulfoxide (−2.22), the values of Landini *et al.*.[31]

TABLE 2. k_{ex}/k_{rac} VALUES OF (+) p-CH_3-C_6H_4-SO-R-^{18}O IN SULFURIC ACID AT 30°C

R;	H_2SO_4 (wt%)	k_{ex}/k_{rac}[c]
CH_3	94.8[b]	0.86
	92.7	0.70
	58.5	0.48
C_6H_5	95.5[b]	0.97
	91.0[b]	0.75
	86.9	0.65
	83.5	0.64
	75.4	0.49[a]

a) Average value of three runs of kinetics at 40, 50 and 60°C. b) At 5°C. c) These rate constants are of pseudo-first-order rate.

TABLE 3. ACTIVATION PARAMETERS OF RACEMIZATION REACTION OF (+) p-$CH_3C_6H_4SO$-R IN SULFURIC ACID

R;	H_2SO_4 (wt%)	E_a (kcal/mol)[a]	$\Delta S^{\neq}$ (e. u.)[a]
CH_3	94.2	20.2	−1.17
	93.8	19.8	−5.94
	64.7	17.1	−25.1
C_6H_5[a]	95.5	24.0	+9.86[b]
	90.8	18.5	−10.8
	85.6	17.5	−17.5
	75.4	18.6	−18.9

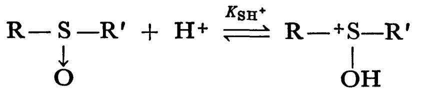
a) These values are based on pseudo first-order rate constants.
b) This is smaller than that reported in the previous report.[10]

route, each cleavage of S–O bond in the rate limiting step which results in oxygen exchange will cause racemization, and $k_{exchange}$ (k_{ex}) is equal to $k_{racemization}$ (k_{rac}), since it passes through a symmetrical intermediate such as $-\overset{+}{S}-$ or $-\overset{++}{S}-$. While, when the reaction proceeds through the S_N2 type route, the k_{ex}/k_{rac} value becomes 0.5.

The k_{ex}/k_{rac} values for I in 94.8% acid and for II in 95.5% acid were found to be 0.86 and 0.97, respectively. While the values for I in 58.5% acid and in 75.4% acid were 0.48 and 0.49, respectively. Thus the predominant path of the reaction in the less concentrated acid is an S_N2 type. And the k_{ex}/k_{rac} value increases gradually with the increase of the acid concentration. However, the values change drastically in above *ca.* 93% acid.

The energies and entropies of activation, E_a and $\Delta S^{\neq}$, obtained in sulfuric acid of several concentrations are listed in Table 3. The E_a value increases significantly while $\Delta S^{\neq}$ value changes from the large negative value to a small negative or positive one with the increase of the acid concentration. In the highly concentrated acid region, the $\Delta S^{\neq}$ values of I and II are -1.77 e.u. (in 94.2% acid) and $+9.86$ e.u. (in 95.5% acid), respectively. While in a less concentrated acid, the $\Delta S^{\neq}$ values are -25.1 e.u. (in 64.7% acid) for I and -18.9 e.u. (in 75.4% acid) for II, respectively. These values are typical for S_N2 type reactions.[11] The drastic change of the $\Delta S^{\neq}$ value is also observed in the higher concentrated acid region above ca. 93—95%.

In these concentrated sulfuric acids, the sulfoxides are considered to be protonated[13] almost completely (Eq. 2). Therefore, the rate can be expressed as a function of the concentration of the protonated sulfoxide, [S⁺–OH], (Eq. 3,4).



$$R-\underset{\underset{O}{\downarrow}}{S}-R' + H^+ \overset{K_{SH^+}}{\rightleftharpoons} R-\underset{\underset{OH}{|}}{{}^+S}-R' \tag{2}$$

$$\text{Rate} = k_1[S^+-OH] \tag{3}$$

$$k_1 = k_{obs}(h_0{}^m + K_{SH^+})/h_0{}^m \tag{4}$$

where, K_{SH^+} and $h_0{}^m$[16] are the thermodynamic ionization constant of the sulfoxides and the acidity function fitted for the sulfoxide-bases, respectively.

In order to solve the rate equation, the accurate pK_{SH^+} values of the sulfoxides are necessary. The pKa values of several sulfoxides have been determined.[17–20] However, it has known that the sulfoxides do not behave as Hammett bases, since the gradient of log [SH⁺]/[S] *vs.* H_0 for the sulfoxides were found to be smaller than unity.[19,20] For example, the average gradient for substituted diphenyl sulfoxides was found to be 0.59.[20b,c] The H_0 scale is based on substituted aniline indicators.[21] Therefore, an acidity function which is strictly appropriate to study especially the ionization behaviors of sulfoxides is most desirable, like the H_A for amide bases,[22,23] and H_0 for benzophenone bases.[24] However, it is difficult to determine the peculiar acidity function just fitted for sulfoxide-bases, and hence it has not been worked out. Yates and McClelland[26] have introduced the H_S function as the ester-base acidity function. And the H_S was obtained by the following empirical relationships, $\log[SH^+]/[S] = -H_s + pK_{SH^+}$; $\log[SH^+]/[S] = -mH_0 +$ const.; $H_S = mH_0$, where m is the slope of plot of the ionization ratio for acetates *vs.* the available H_0 scale. Here, we also undertook the same treatment for the sulfoxide-base. From the average gradient of log [SH⁺]/[S] vs H_0 measured by UV and NMR, mentioned above, we have chosen to use m=0.60 and 0.6 H_0 instead of H_0 in this study.[27] Recently, Landini *et al.*[31,33] have found that the H_A function developed by Yates *et al.*[22] satisfactorily represents the protonation behaviors of the sulfoxides. The H_A is about 0.65 H_0 (between 50 and 80% acid), which also corresponds closely to the approximation, $m=0.6$, for the value of sulfoxide-bases mentioned above. Then the rate constants, k_1, of both the oxygen exchange and racemization reactions were calculated by the eq. 4 using h_A and $h_0{}^{0.6}$ and the pK_{SH^+} values, where $h_0{}^{0.6}$ was used to conduct the kinetic data in higher than 80% acid.

The rate constants, k_1, obtained are listed in Table 1, which also increase with the increase of the acid concentration. The acid dependencies of the rates are collected in Table 4. At first, diagrams correlating the logarithmus of the rates of racemization of both I and II against H_A(or $0.6H_0$) were made. At highly concentrated acid regions, log k_1 was found to be in a

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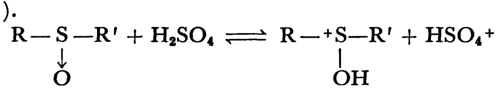
[11] For example, the $\Delta S^{\neq}$ value for the racemization reaction of II with acetic anhydride which is a typical S_N2 type process was found to be -28.6 e. u.[12]

[12] S. Oae and M. Kise, *Tetrahedron Lett.*, **1967**, 1409.

[13] It has been shown by Gillespie and Passerini[14] and Oae *et al.*[15] through the cryoscopic measurement that the sulfoxides ionize in the concentrated sulfuric acid in the following manner (eq. 1).



$$R-\underset{\underset{O}{\downarrow}}{S}-R' + H_2SO_4 \rightleftharpoons R-\underset{\underset{OH}{|}}{{}^+S}-R' + HSO_4{}^+ \tag{1}$$

[14] R. J. Gillespie and R. C. Passerini, *J. Chem. Soc.*, **1956**, 3850.

[15] S. Oae, T. Kitao, and Y. Kitaoka, This Bulletin, **38**, 543 (1965).

[16] This symbol is one given by Yates in his paper.[25]

[17] P. Nylen, *Z. Anorg. Allgem. Chem.*, **246**, 227 (1941).

[18] K. K. Andersen, W. H. Edmonds, J. B. Biasotti, and R. A. Strecker, *J. Org. Chem.*, **31**, 2859 (1966).

[19] a) P. Haake and R. D. Cook, *Tetrahedron Lett.*, **1968**, 427. b) P. Haake, R. D. Cook, and G. H. Hurst, *J. Amer. Chem. Soc.*, **89**, 2650 (1967).

[20] a) S. Oae, K. Sakai, and N. Kunieda, This Bulletin, **42**, 1966 (1969). b) N. Kunieda and S. Oae, unpublished data. c) Haake and Cook also reported the similar small slopes for phenyl methyl sulfoxide (0.62) and dimethyl sulfoxide (0.61) in the pK_{SH^+} determinations by NMR chemical shift measurements of methyl group in sulfuric acid.[19]

[21] L. P. Hammett, "Physical Organic Chemistry", McGraw-Hill Book Co. Inc., New York, N. Y., 1970, Chap. 9.

[22] K. Yates, J. B. Stevens, and A. R. Katritzky, *Can. J. Chem.*, **42**, 1957 (1964).

[23] K. Yates and J. B. Stevens, *ibid.*, **43**, 529 (1965).

[24] T. G. Bonner and J. Phillips, *J. Chem. Soc.* (B), **1966**, 650.

[25] K. Yates, *Accounts Chem. Res.*, **4**, 136 (1971).

[26] K. Yates and R. A. McClelland, *J. Amer. Chem. Soc.*, **89**, 2686 (1967).

[27] The H_0 values used here are the data reevaluated more recently by Johnson *et al.*[30] The measurements of ionization ratio of sulfoxides by us[20] and Haake and Cook[19] refer to H_0 scales compiled by Long and Paul[28] and Jorgenson and Hartter,[29] which will not be strictly correct when the Johnson *et al.* values are used. However, since the difference of the two H_0 scales is at most 0.2 H_0 at all acid range, it may be parmitted to use these values of ionization ratios.

[28] M. A. Paul and F. A. Long, *Chem. Revs.*, **57**, 1 (1957).

[29] M. J. Jorgenson and D. R. Hartter, *J. Amer. Chem. Soc.*, **85**, 878 (1963).

[30] C. D. Johnson, A. R. Katritzky, and S. A. Shapiro, *ibid.*, **91**, 6654 (1969).

[31] D. Landini, G. Modena, G. Scorrano, and F. Taddei, *ibid.*, **91**, 6704 (1969).

[32] J. F. Bunnett and F. P. Olsen, *Can. J. Chem.*, **44**, 1899 (1966).

[33] D. Landini, G. Modena, F. Montanari, and G. Scorrano, *J. Amer. Chem. Soc.*, **92**, 7168 (1970).

TABLE 4. CORRELATION BETWEEN ACIDITY FUNCTIONS AND RATE CONSTANTS OF RACEMIZATION OF p-CH_3-C_6H_4SOR IN SULFURIC ACID

R;	H_2SO_4 (wt%)[a]	log k_1 vs. H_A ($0.6H_0$)	log $k_{obs}+H_0$ vs. log a_{H_2O}	log $k_{obs}+H_A$ ($0.6H_0$) vs. log a_{H_2O}	log k_1+H_0 vs. H_0+log [H^+]
CH_3	52.4—68.8	0.583	1.86 (0.999)	0.388 (0.988)	0.770 (0.999)
	89.8—94.2	1.18 (0.997)[b]	0.328 (0.985)	−0.116[c]	0.336 (0.986)
C_6H_4	64.7—74.4	0.750	1.44 (0.999)	0.317 (0.949)	0.701 (0.991)
	93.3—96.3	1.44 (0.990)	0.133[c]	−0.272 (0.866)	0.300 (0.880)

a) Range of sulfuric acid concentration. b) Correlation coefficient.
c) It's correlation showed as low value because the slope is very low.

linear relationship with $0.6H_0$. The slopes for I at above 89.8% acid and for II at above 93.3% acid were 1.18 and 1.44, respectively. This means that the reaction proceeds through a predominant A-1 type route, according to the Zucker-Hammett hypothesis.[34] As the concentration of acid decreases, the plot of log k_1 *vs.* acidity functions start to deviate markedly from the linear slopes to lower ones.

The Bunnett *w*-values,[35] at the range of less concentrated acid for I(below 68.8%) and for II(below 74.4%) are 1.86 and 1.44, respectively. Whereas, at sulfuric acids of above 89.9% for I and above 93.3% for II, the *w*-values are 0.328 and 0.133, respectively. Though the values are little larger than zero and cannot be classified strictly into the class $w \leq 0$ of the criterion, the values seem to indicate that the reaction proceeds through a nearly A–1 type route in which water does not participate in the rate limiting step. According to the modified Bunnett hydration parameter treatment[23,25,36] using H_A or $0.6H_0$ for H_X, the slopes (r), at the highly concentrated acid region, show smaller values than zero (*i. e.* for I; −0.116 (above 89.8%), for II; −0.272 (above 93.3%)), and at the less concentrated acid region, however, the slopes show small positive values. However, the equation, $\log k_{obs} - \log h_A \text{ (or } h_0^{0.6})/[K_{SH^+} + h_A \text{ (or } h_0^{0.6})] = r \log a_{H_2O} +$ const., gives a large negative r values at all range of the plot both for I and II, even at the less concentrated acid region.

The plot[32] of ($\log k_{obs}+H_0$) against (H_0+log [H^+]) also gave a linear slope at the less concentrated acid range. The ϕ values for I and II are 0.770 and 0.701, respectively. On the other hand, at the range of concentrated acid, the slopes for I, and II were larger than that expected from a typical A–1 mechanism. The plot of ($\log k_1+H_0$) *vs.* (H_0+log [H^+]) showed a similar slope.[37] These results seem to indicate that the S_N2 type reaction takes place competitively in no small portion, even in the high acid region above 85%. The plot of $\log k_{obs} - \log h_0^m/(K_{SH^+}+h_0^m)$ *vs.* (H_0+log [H^+]), where h_0^m is h_A or $h_0^{0.6}$, showed high negative ϕ values for both I and II at all ranges of the acid concentration, even including less concentrated acids. Similar large negative ϕ values have been observed, for instance, in the reduction of sulfoxides by halide ion(−1.2∼−1.6)[33] and Wallach rearrangement(−1.8 ∼−0.4),[38] which have been believed to involve two protons in the rate limiting step.

It would be difficult to apply the Bunnett[35] and the Bunnett and Olsen[32] criteria to analyze directly the reaction at the high acid region.

TABLE 5. EFFECT OF DEUTERIOSULFURIC ACID ON THE RACEMIZATION RATE OF I AND II

Sulfoxide	Sulfuric acid (mol/l)	Temp. (°C)	$\times 10^4\ s^{-1}$ [a]		k_H/k_D
			$k_{H_2SO_4}$	$k_{D_2SO_4}$	
I	17.48	10	18.6	13.9	1.34
	17.09	10	10.8	8.71	1.24
	14.61	25	3.10	2.60	1.19
	11.40	45	2.43	2.57	0.946
	10.27	30	0.306	0.322	0.950
	10.27	45	1.48	1.59	0.931
II	17.68	10	7.00	6.93	1.01
	13.02	45	1.97	1.78	1.11

a) Pseudo first order rate constants using Method B. These values are average ones for three or four measurements.

Table 5 summarizes the values of the ratio of the rate, k_H/k_D, *i. e.* the ratio of the rates of the racemization of I and II in D_2SO_4–D_2O media to those in ordinary H_2SO_4–H_2O media. The solvent isotope effect observed is not very large. As described earlier, in both highly and moderately concentrated sulfuric acids, the sulfoxides are considered to be protonated almost completely. Therefore, if the reaction proceeds through an A–1 type route, the isotope effect would be found only in the step of unimolecular S–O bond cleavage of the protonated sulfoxides, and its k_H/k_D value at the concentrated acid regions should be close to unity. In the case of II, the value is indeed 1.01 at the acid of 17.68 mol/l. However, the value of I was found to be a larger than one, *i.e.* 1.34, in the acid of similar concentration, *i.e.* 17.48 mol/l. This means that, in the case of I, an S_N2 type reaction involving the nucleophilic attack of water on the protonated sulfoxide takes place to substantial extent even at the high concentrated acid region, and hence the solvent

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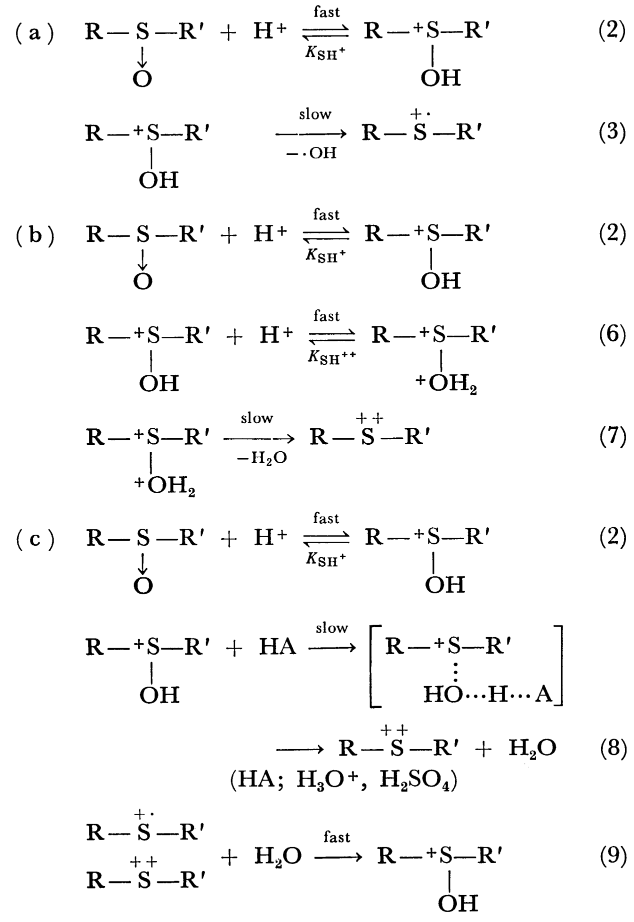
[34] M. A. Paul and F. A. Long, *Chem. Revs.*, **57**, 935 (1957); L. Zucker and L. P. Hammett, *J. Amer. Chem. Soc.*, **61**, 2791 (1939).

[35] J. F. Bunnett, *ibid.*, **82**, 499 (1960); **83**, 4956 (1961); **83**, 4968 (1961); **83**, 4973 (1961); **83**, 4978 (1961).

[36] K. Yates and J. C. Riordan, *Can. J. Chem.*, **43**, 2328 (1965).

[37] Landini *et al.* applied the same treatment using k_1 into a study of the reduction of sulfoxides by halide ion in aqueous perchloric acid, and they gave the ϕ values in the range from −0.12 to +0.06.[33]

[38] E. Buncel and B. T. Lawton, *Can. J. Chem.*, **43**, 862 (1965).



Scheme 1. A-1 like mechanism.

isotope effect due to the difference of the nucleophilicities between D_2O and H_2O would appear. As the acid concentration decreases, the k_H/k_D value of I also decreases gradually, and at 10.27 mol/l, the isotope effect is reversed, *i.e.* $k_H/k_D=0.931$. Thus it is clear that the initial protonation equilibrium controls the isotope effect, because in the less concentrated acids the protonation would not be complete.

There are a few concievable pathways for the A-1 type S-O bond cleavage that fit to these observations, as shown in Scheme 1.

The several ESR studies of the sulfur containing aromatic compounds have been done.[39] Shine *et al.*[40] revealed that noticeable ESR signals of the corresponding cation radical were observed from 96% sulfuric acid solutions of *p,p'*-disubstituted diphenyl sulfoxides. Recently, we have also found the ESR signals from 96.3% sulfuric acid solutions of *p,p'*-ditolyl sulfoxide and II,[9] though the signal from II has no hyperfine structure. These ESR results suggest that the route (a) involving the homolytic cleavage to form an incipient cation radical intermediate is the most likely possibility, in the case of diaryl sulfoxides. The polar effects of *p*-substituents on the oxygen exchange reaction rates are very small in comparison with that of the typical S_N1 type reactions (see Table 6). This rather small polar effect also seems to favor the route (a). While the 96% sulfuric acid solution of I is colorless,[41] and no noticeable ESR signal was detected for I from the 96% acid solution. It is dubious, therefore, that I undergoes the reaction depicted by route (a), in the acid below 96% concentration.

Table 6. Rate constants of ^{18}O-exchange of p-X-$C_6H_4S^{18}O$-C_6H_5[a] in 95.3% sulfuric acid at 5°C

X;	$10^4 k_{ex}$ (sec^{-1})[b]
CH_3	2.92±0.20
H	1.15±0.08
Cl	1.53±0.05
NO_2	1.13±0.03

a) Sulfoxide=0.12 mol/l, by Method A.
b) These values are much the same as those of our previous work.[9]

The pathways (b) and (c), including the dication intermediate, must also be included among the A-1 type reactions. For the unimolecular S-O bond cleavage, the second protonation or a second proton transfer from acid media to the mono-protonated sulfoxide is considered to be neccessary, since, OH group being a poor leaving group, the S^+-OH should pick up one more proton before the heterolytic cleavage of S-O bond. Route (b) begins with a rapid protonation equilibrium (Eq. 2) and followed by a second protonation (Eq. 7) again on the oxygen. Then the rate-limiting step becomes the loss of a water molecule from the diprotonated sulfoxide. This dication intermediate is similar to the carbonium ion in the acid catalyzed oxygen exchange reaction of alcohols.[42]

If the reaction proceeds through the overall two-proton process, the rate is expressed by the use of the H_+ function.[43] The kinetic analyses with the H_+ function also permit the reaction in the concentrated acid to follow the route (b). Though we can find no value of H_+ available for this treatment, the parallelism has been found between H_+ and H_0 in the sulfuric acid media,[43-44] and hence H_A or $0.6H_0$ could be substituted for H_+ in this instance,[45] even though

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[39] U. Schmidt, K. Kabitze and K. Markau, *Angew. Chem.*, **72**, 708 (1960); U. Schmidt, *ibid.*, **76**, 629 (1964); H. J. Shine, "The Formation of Cations and Cation Radicals from Aromatic Sulfides and Sulfoxides," in "Organosulfur Chemistry," ed by M. J. Janssen, Intra-Science (1967), Chap. 6. and the references cited therin.

[40] H. J. Shine, M. Rahman, H. Seeger and G.-S. Wu, *J. Org. Chem.*, **32**, 1901 (1967).

[41] The color of the solution of I at nearly 100% acid is orange.

[42] The references cited in; D. Samuel and B. L. Silver, "Oxygen Isotope Exchange Reactions of Organic Compounds", Vol. 3, in "Advances in Physical Organic Chemistry", ed. by V. Gold, Academic Press, New York, N. Y. (1965), Chap. 3.

[43] T. G. Bonner and J. C. Lockhart, *J. Chem. Soc.*, **1957**, 364; J. C. D. Brand, W. C. Horning, and M. B. Thornley, *ibid.*, **1952**, 1374; P. Vetešnik, J. Bielavský, and M. Večeřa, *Collect. Czeck. Chem. Commun.*, **33**, 1687 (1968).

[44] M. Isaks and H. H. Jaffé, *J. Amer. Chem. Soc.*, **86**, 2210 (1964).

[45] This approximation has been reported to be applied satisfactorily for the mechanistic interpretation of the reactions involving the two protonation equilibrium such as hydrolysis of 4-dimethoxymethylpyridinium ion[46] and Wallach rearrangement.[38]

[46] M. W. Fuller and W. M. Schubert, *J. Amer. Chem. Soc.*, **85**, 108 (1963).

Table 7. Rates of racemization of sulfoxides *p*-Tolyl-SO-R

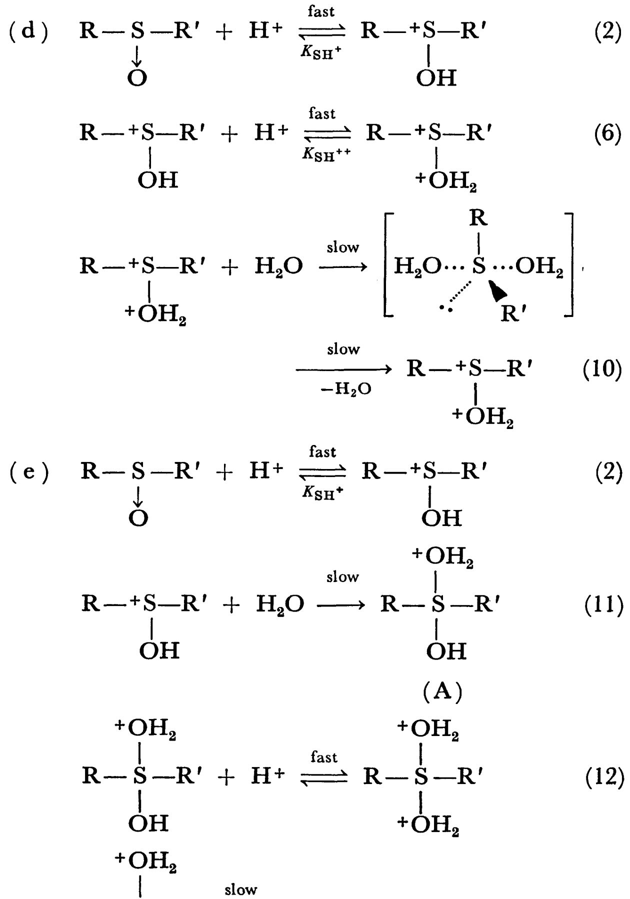
R;	64.7% H_2SO_4 a) 10^4k (s^{-1})	96.7% H_2SO_4 b) 10^4k (s^{-1})	HCl c) 10^5k (s^{-1})	HCl d) 10^6k (s^{-1})
CH_3	4.20 ±0.40(3.53)	7.58 ±0.25(3.18)	3700(463)	
C_2H_5	2.30 ±0.15(1.93)		957(120)	
Phenyl	1.19 ±0.03(1.00)	2.38 ±0.10(1.00)	8.00(1.00)	27.6(1.00)
o-Tolyl	0.557±0.018(0.468)		4.53(0.566)	
Mesityl		0.268±0.013(0.113)	0.042(0.0052)	0.125(0.0045)

a) Sulfoxide=0.16 mol/l, at 30 °C. b) Sulfoxide=0.1 mol/l, at 0 °C.
c) In a 1 : 2 v/v mixture of 12 M aqueous HCl and dioxane at 25 °C (Ref. 1).
d) In 3.93 M-HCl in 80% dioxane-water, at 25 °C (Ref. 2). e) The relative rates are given in parentheses.

an error may be introduced.

Meanwhile, in the strong acid above *ca.* 93%, all acid species in the medium must take the form of either H_3O^+ or H_2SO_4. Then the route (c) including the general acid catalyzed removal of water would be conceivable. However, according to this mechanism, since it includes a proton transfer from the acid medium to the mono-protonated sulfoxide, a large isotope effect, $k_H/k_D>1$, must be expected. Actually, the isotope effect observed in the racemization of I in 17.68 mol/l is nearly one, as shown in Table 5. So this route (c) can be excluded.

All these results seem to suggest that sulfoxide I undergoes much less the A-1 reaction than II, and the legitimate A-1 type reaction takes place with I only in an extremely concentrated sulfuric acid close to 100%. This may be attributed to the higher stability of the monoprotonated sulfoxide I and the less stability of the intermediate such as cation radical or dication derived from I.



$$\text{(d)}\quad R\text{—}\underset{\substack{\downarrow\\ O}}{S}\text{—}R' + H^+ \underset{K_{SH^+}}{\overset{\text{fast}}{\rightleftharpoons}} R\text{—}\underset{\substack{|\\ OH}}{{}^+S}\text{—}R' \qquad (2)$$

$$R\text{—}\underset{\substack{|\\ OH}}{{}^+S}\text{—}R' + H^+ \underset{K_{SH^{++}}}{\overset{\text{fast}}{\rightleftharpoons}} R\text{—}\underset{\substack{|\\ {}^+OH_2}}{{}^+S}\text{—}R' \qquad (6)$$

$$R\text{—}\underset{\substack{|\\ {}^+OH_2}}{{}^+S}\text{—}R' + H_2O \xrightarrow{\text{slow}} \left[H_2O\cdots\overset{\substack{R\\|}}{S}\cdots OH_2\right] \xrightarrow[-H_2O]{\text{slow}} R\text{—}\underset{\substack{|\\ {}^+OH_2}}{{}^+S}\text{—}R' \qquad (10)$$

$$\text{(e)}\quad R\text{—}\underset{\substack{\downarrow\\ O}}{S}\text{—}R' + H^+ \underset{K_{SH^+}}{\overset{\text{fast}}{\rightleftharpoons}} R\text{—}\underset{\substack{|\\ OH}}{{}^+S}\text{—}R' \qquad (2)$$

$$R\text{—}\underset{\substack{|\\ OH}}{{}^+S}\text{—}R' + H_2O \xrightarrow{\text{slow}} R\text{—}\overset{\substack{{}^+OH_2\\|}}{\underset{\substack{|\\ OH}}{S}}\text{—}R' \;\text{(A)} \qquad (11)$$

$$R\text{—}\overset{\substack{{}^+OH_2\\|}}{\underset{\substack{|\\ OH}}{S}}\text{—}R' + H^+ \overset{\text{fast}}{\rightleftharpoons} R\text{—}\overset{\substack{{}^+OH_2\\|}}{\underset{\substack{|\\ {}^+OH_2}}{S}}\text{—}R' \qquad (12)$$

$$R\text{—}\overset{\substack{{}^+OH_2\\|}}{\underset{\substack{|\\ {}^+OH_2}}{S}}\text{—}R' \xrightarrow{\text{slow}} R\text{—}\underset{\substack{|\\ {}^+OH_2}}{{}^+S}\text{—}R' + H_2O \qquad (13)$$

Scheme 2. *S*N2-like mechanism.

The following two alternative routes (d) and (e) (in Scheme II) may be postulated for the S_N2 type reaction.

There are differences between the two routes (d and e) in the second protonation step and in the degree of bonding of both entering and leaving groups on sulfur atom at the transition state. The reaction *via* route (d) includes the rate-limiting displacement of water molecules on the sulfur atom. And the every displacements cause the "Walden Inversion" at the sulfur atom, since both entering and leaving water molecules assume axial-axial positions at the trigonal bipyramidal transition.[47] The scheme *via* route (e) begins with the first equilibrium of mono-protonation, followed by the slow nucleophilic attack of water to form intermediate (A), and the subsequent fast second protonation to (A). The racemization may take place in the step 11 or 12 or both, and the oxygen exchange takes place at the step 13. Up to now, many reactions including tetracovalent pentacoordinate sulfur intermediates have been investigated. The acid catalyzed reduction of sulfoxides by halide ion[33,48] and recemization reaction of sulfoxides in aqueous hydrochloric acid[1,2] have been known to form a halohydrine intermediate ($R_1R_2S(OH)X$), and in these reactions, large steric effects on the rates by bulky substituents attacked to the central sulfur atom have been found owing to the tightness of the intermediate. If the title reaction would take place through the route (e), a similarly large steric effect must be expected. The steric effects observed in the racemizations of *p*-tolyl alkyl and *p*-tolyl aryl sulfoxides in 64.7% sulfuric acid were studied. However, as shown in Table 7, the steric effect is not so large and almost the same as that in 96.7% sulfuric acid. Thus, the route (e) including the intervention of a tight intermediate seems unlikely.

The rate constants, k_1, of the oxygen exchange of several *p*-substituted diphenyl and methyl phenyl sulfoxides were measured in the less concentrated acid.

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[47] S. Oae, *Quart. Rept. Sulfur Chem.*, **5**, 53 (1970).

[48] J. H. Krueger, *Inorg. Chem.*, **5**, 132 (1966); D. Landini, F. Montanari, G. Modena, and G. Scorrano, *Chem. Commun.*, **1968**, 86; R. A. Strecker and K. K. Andersen, *J. Org. Chem.*, **33**, 2234 (1968).

TABLE 8. RATE CONSTANTS OF ^{18}O-EXCHANGE OF SUBSTITUTED SULFOXIDES R_1-$S^{18}O$-C_6H_4-R_2

R_1;	R_2;	H_2SO_4 (wt%)	10^4k_{obs} (s^{-1})	10^4k_1 (s^{-1})c)
Phenyla)	p-CH_3	75.4	1.45	1.54
Phenyl	H	75.4	1.85	2.08
Phenyl	p-Cl	75.4	1.98	2.30
Phenyl	p-NO_2	75.4	2.43	
CH_3b)	p-CH_3	64.7	0.202	0.216
CH_3	H	64.7	0.238	0.257
CH_3	p-Cl	64.7	0.241	0.270

a) Sulfoxide=0.1 mol/l, at 50°C,10) Method A.
b) Sulfoxide=0.1 mol/l, at 30°C, Method A.
c) k_1 was calculated using the pK_{BH^+} values established by Landini *et al.*31) (for methyl phenyl sulfoxides), and by us (for diphenyl sulfoxides).

One finds in Table 8 that the polar effects of *p*-substituents on k_1 are not large. When log k_1 values obtained were plotted against σ^+, a good straight line was obtained (for diphenyl sulfoxides $\rho=0.412$, and for methyl phenyl sulfoxides $\rho=0.250$). However, the small positive values of ρ seems only to point out that the nucleophilic substitution of the protonated sulfoxide by water is the slow rate-limiting step of the reaction at a less concentrated sulfuric acid.

Experimental

Materials. *Optically Active Sulfoxides* were prepared from (—)menthyl (—) *p*-toluenesulfinate [mp 107—7.5 °C, $[\alpha]_D$ —200.4° (c=1.85, l=0.5 acetone)(lit,49) mp 106—7 °C, $[\alpha]_D^{25}$ —199.2° (acetone)] and the corresponding aryl or alkyl Grignard's reagent according to the method developed by Andersen.50) The mp or bp and specific rotations of sulfoxides obtained are as follows. *p*-Tolyl-SO-R; (R; mp or bp, specific rotation (l=0.5), (lit)51)). CH_3; 74 °C, $[\alpha]_D^{23}$ +146.0° (c=1.80, acetone), (73—74.5 °C, $[\alpha]_D$ +145.5° (acetone)), C_2H_5; 95—97 °C/0.5 mmHg, $[\alpha]_D^{25}$ +184° (c=1.3, acetone), (123—126 °C/1.5 mmHg, $[\alpha]_D^{25}$ +186°(acetone)), C_6H_5; 92 °C, $[\alpha]_D^{20}$ +27.6°(c=1.53, EtOH), (92—93 °C, $[\alpha]_D$ +27°(EtOH)), *o*-CH_3-C_6H_4; 78—78.5 °C, $[\alpha]_D^{20}$ —90.5°(c=1.21, acetone), (77.8—78.5 °C, $[\alpha]_D$ —89.1° (acetone)), Mesityl; 107—107.5 °C, $[\alpha]_D^{22}$ —265.5° (c=1.53, EtOH), (107—107.5 °C, $[\alpha]_D$ —259° (EtOH)).

^{18}*O-Labeled Sulfoxides* were prepared by the bromine-oxidation of the corresponding sulfides in the presence of ^{18}O-enriched water(1.62 *atom*%-^{18}O) in acetic acid by the similar procedure described in our previous paper.52) A sulfide (1 mol) was dissolved in a mixture of ^{18}O-enriched water(2—5 mol, 1.62 *atom*%-^{18}O), pyridine(5 mol) and acetic acid, and the solution was cooled under 0 °C, and then acetic acid solution of bromine (1 mol) was added dropwise over a period of 1 hr. After stirring for 1 hr, the mixture was poured into a large amount of ice-water, and the aqueous solution was neutralized with sodium bicarbonate. The resulted sulfoxide was extracted with chloroform, and the extract was washed with 10% aqueous solution of sodium thiosulfate and water, respectively, and dried. Finally, chloroform was distilled off and the crude ^{18}O-labeled sulfoxide was purified by distillation or recrystallization. Their mp or bp and *atom*%-^{18}O are as follows. $C_6H_5SOC_6H_4$-R; (R; mp or bp(lit)53), *atom*%-^{18}O), *p*-CH_3; 71.5—72 °C (71 °C), 1.12, H; 70—71 °C(71 °C), 0.852, *p*-Cl; 45—46 °C (45—46 °C), 1.23, *p*-NO_2; 107 °C (106—107 °C), 0.996(This oxide was prepared using 3.50 *atom*%-$H_2{}^{18}O$). $CH_3SOC_6H_4$-R; (R; mp or bp(lit), *atom*%-^{18}O), *p*-CH_3; 42—43 °C (42—43 °C),53) 1.08, H; 84 °C/0.25 mmHg(85 °C/0.5 mmHg),54) 0.997, *p*-Cl; 46—48 °C(47—48 °C),53) 0.983.

Kinetic Procedures. The rate of oxygen exchange and racemization were measured by the following two methods using YANAGIMOTO OR-10 TYPE polarimeter and HITACHI RMU-6E TYPE mass-spectrometer.

Method A (for both oxygen exchange and racemization). An ^{18}O-labeled or an optically active sulfoxide of a set mol was dissolved in sulfuric acid at a preset temperature. Aliquot portion(5 ml) of the solution was taken up from time to time for 5—7 times, and was quenched with 100 ml of ice-water. The sulfoxide recovered was then extracted with carbon tetrachloride or chloroform and the extract was washed with water, 10% aqueous solution of sodium carbonate and water, respectively, and dried over anhydrous sodium sulfate, and the solvent was then distilled off. The sulfoxide resulted was dried and subjected to both ^{18}O-analysis or specific rotation measurement. The ^{18}O-atom% was calculated from the mass peak heights 44 and 46 of carbon dioxide according to the method described by Samuel.55) The rate constant for oxygen exchange were calculated from the equation, $\log \beta_0-\beta/\beta_t-\beta=kt/2.303$, where β_0 and β_t are atom% of ^{18}O at time 0 and t, respectively, and β is the atom% of ^{18}O of natural CO_2. The pseudo-first-order rate constant for the racemization was calculated from the equation, $\log \alpha_0'/\alpha_t'=kt/2.303$, where α_0' and α_t' are the specific rotation, $[\alpha]_D$, at time 0 and t, respectively.

Method B (for racemization): In the square shape cell (5 cm) was placed a sulfuric acid solution containing an optically active sulfoxide of a set mol, the rate was measured directly by checking the rotation, α, of polarimeter which was set at a desired temperature with a constant temperature bath connected. The pseudo-first-order rate constants were evaluated by the following equation, $\log \alpha_0/\alpha_t=kt/2.303$, where α_0 and α_t are the rotatory power at time 0 and t, respectively.

Sulfuric Acid. The concentration(wt%) of sulfuric acid medium was determined by titration with a standard alkali solution. The H_0 values of the media were obtained by interpolation from a graph of the data of Johnson *et al.*30) D_2SO_4-D_2O solutions were prepared from MERCK D_2SO_4 (*min.* 99%-*d*) by dilution with D_2O (*min.* 99.75%-*d*). The concentration of acid was determined by titration of a weighed sample with a standard alkali solution. The molarity of the acid was calculated from the specific gravity. The specific gravities of ordinary acids used here are the values in the established table,56) and these of deuterio acid media were newly determined by weight-method. The activity of water, log a_{H_2O}, of the acid media was the value interpolated from the table of Giauque *et al.*57)

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49) H. F. Herbrandson, R. T. Dickerson, Jr., and J. Winstein, *J. Amer. Chem. Soc.*, **78**, 2576 (1956).

50) K. K. Andersen, *Tetrahedron Lett.*, **1962**, 93.

51) K. Mislow, M. M. Green, P. Laur, J. T. Melillo, T. Simmons, and A. L. Ternay, Jr., *J. Amer. Chem. Soc.*, **87**, 1958 (1965).

52) S. Oae, Y. Ohnishi, S. Kozuka, and W. Tagaki, This Bulletin, **39**, 364 (1966).

53) A. Cerniani and G. Modena, *Gazz. Chim. Ital.*, **89**, 834 (1959).

54) K. K. Andersen, *J. Org. Chem.*, **29**, 1953 (1965).

55) D. Samuel, "Methodology by Oxygen Isotope," in "Oxygenease," Chap. 2, ed by O. Hayaishi, Academic Press, New York, N. Y., (1962).

56) "Kagaku Binran", ed by Chem. Soc. Japan. Maruzen, Japan (1966), p. 439.

57) W. F. Giauque, E. W. Hornung, J. E. Kunzler, and T. R. Rubin, *J. Amer. Chem. Soc.*, **82**, 62 (1960).