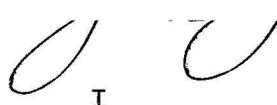


AN ABSTRACT OF THE THESIS OF

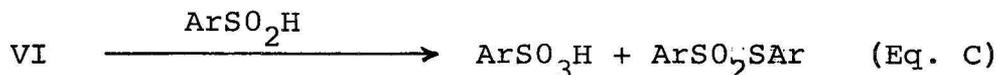
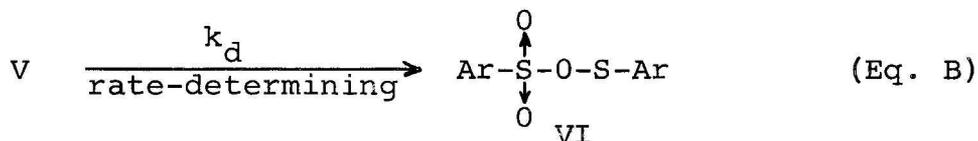
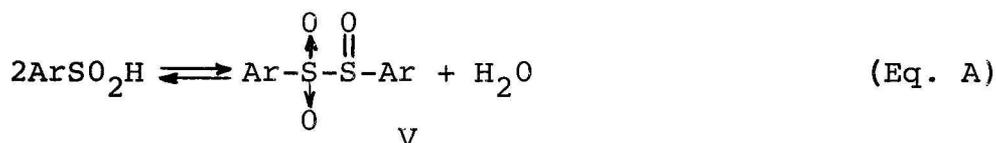
Alvin Fitzgerald for the M.S. in Chemistry
 (Name) (Degree) (Major)

Date thesis is presented July 13, 1965

Title Studies of the Mechanism of the Disproportionations
of Sulfinic Acids and Disulfides

Abstract approved Redacted for Privacy


Previous studies of the mechanism of the disproportionation of arylsulfinic acids have provided quite convincing evidence that the mechanism of the reaction is as follows:



In the present work the effect of substituents on the rate of disproportionation of a series of para-substituted benzenesulfinic acids has been determined. The rates of disproportionation were found to be

correlated by σ^+ substituent constants. The effect of substituent on the rate was shown to be mainly due to its effect on the position of the first equilibrium.

Data on the rate of disproportionation of p-toluenesulfinic acid as a function of temperature allow one to calculate the enthalpy of activation for the reaction.

Comparison of both rate and enthalpy of activation for the disproportionation with the corresponding quantities previously determined in dioxane solution for the decomposition of p-toluenesulfinyl p-tolyl sulfone (Equation B) make it possible to plot, at least qualitatively, free energy-reaction profile diagrams for the disproportionation of arylsulfinic acids.

II

The disproportionation or interchange reaction of benzyl disulfide-p-tolyl disulfide (Equation D) has been



studied in acetic acid-sulfuric acid-water solution using a nuclear magnetic resonance spectroscopy method.

The equilibrium constant for this reaction is somewhat less than the value ($K_e=4$) predicted on purely statistical grounds. This is the first case of a disulfide interchange where K_e has been found to be less than 4.

Some preliminary studies of the rate of equilibration have also been carried out. The reaction was found to be catalyzed by strong acid. However, the data, when plotted as a second-order equilibrium reaction, gave non-linear plots. Reasons for this behavior are discussed.

STUDIES OF THE MECHANISM OF THE
DISPROPORTIONATIONS OF SULFINIC ACIDS AND DISULFIDES

by

ALVIN FITZGERALD

A THESIS

submitted to

OREGON STATE UNIVERSITY

in partial fulfillment of
the requirements for the
degree of

MASTER OF SCIENCE

June 1966

APPROVED: Redacted for Privacy

Associate Professor of Chemistry
In Charge of Major

Redacted for Privacy

Chairman of Department of Chemistry

Redacted for Privacy

Dean of Graduate School

Date thesis is presented July 13, 1965

Typed by Erma McClanathan

ACKNOWLEDGMENT

The author wishes to thank Dr. John L. Kice for his help and advice during the course of this work. Thanks are also expressed to Dr. Mary Dankleff for determining the nuclear magnetic resonance spectra.

TABLE OF CONTENTS

PART I: THE DISPROPORTIONATION OF SULFINIC ACIDS

INTRODUCTION.....	1
The Structure of Sulfinic Acids.....	1
Chemical Reactions of Sulfinic Acids.....	2
Disproportionation of Sulfinic Acids.....	3
RESULTS.....	13
Kinetics of Disproportionation.....	13
Effect of Structure on Rate of Disproportionation	14
Determination of Activation Parameters.....	17
DISCUSSION.....	20
Thermodynamic Activation Parameters.....	20
Effect of Structure on Sulfinic Acid Disproportionation.....	25
EXPERIMENTAL.....	28
1. Preparation of Sulfinic Acids.....	28
p-Bromobenzenesulfinic Acid.....	28
p-Methoxybenzenesulfinic Acid.....	28
2-Naphthalenesulfinic Acid.....	29
p-Nitrobenzenesulfinic Acid.....	29
p-Toluenesulfinic Acid.....	30
2. Preparation of Stock Solutions.....	30
Glacial Acetic Acid.....	30
3. Procedures for Following Kinetic Experiments with Substituted Arylsulfinic Acids.....	32
General Method for a Kinetic Run.....	32
Kinetic Studies of p-Methoxybenzenesulfinic Acid.....	35
Kinetic Procedure for p-Nitrobenzenesulfinic acid.....	35

PART II: THE DISPROPORTIONATION OF DISULFIDES

INTRODUCTION.....	37
Methods of Preparation.....	37
Reactions of Disulfides.....	39
Disulfide Disproportionation or Interchange Reaction.....	40
RESULTS.....	48
Equilibrium Studies on the Benzyl Disulfide - p-Tolyl Disulfide Interchange System.....	48
Rate of Equilibration Measurements.....	51
Dependence of Rate on Acidity.....	52
DISCUSSION.....	55
Equilibrium Constants for the Benzyl Disulfide- p-Tolyl Disulfide Interchange.....	55
Acidity Dependence of the Disulfide Disproportionation.....	56
Curvature of Kinetic Plots.....	57
EXPERIMENTAL.....	59
1. Purification of Symmetrical Disulfides.....	59
Benzyl Disulfide.....	59
p-Tolyl Disulfide.....	59
2. Preparation of Stock Solutions.....	59
3. Procedure for Determining Equilibrium Constants.....	60
4. Procedure for Following the Rate of the Benzyl Disulfide p-Tolyl Disulfide Interchange.....	61
BIBLIOGRAPHY.....	63
APPENDIX I.....	71
APPENDIX II.....	76

LIST OF TABLES

Table		
I	Thermodynamic Quantities for Sulfinyl Sulfone Decomposition.....	7
II	Effect of Structure on the Rate of Sulfinyl Sulfone Decomposition.....	8
III	Ease of SO ₂ Loss of Some Sulfinic Acids on Heating in Solution.....	9
IV	Effect of Structure on Sulfinic Acid Disproportionation.....	9
V	Effect of Sulfuric Acid Concentration on Rate of Disproportionation.....	10
VI	Effect of Water on Rate of Disproportionation	11
VII	Kinetics of the Disproportionation of Aromatic Sulfinic Acids in Acetic Acid.....	14
VIII	Effect of Structure on Rate of Disproportionation of Aromatic Sulfinic Acids.....	15
IX	Ratios of k ₂ at 80°C to k ₂ at 70°C for Varying Sulfuric Acid Concentration and Aryl Group.....	18
X	Equilibrium Constants for Some Dialkyl Disulfide Interchange Reactions.....	44
XI	Equilibrium Constants for Some Dialkyl Disulfide Interchange Reactions.....	45
XII	Equilibrium Data for the Benzyl Disulfide-p-Tolyl Disulfide Interchange.....	50
XIII	Relationship of H ₀ to Log t ₀ for the Benzyl Disulfide-p-Tolyl Disulfide Interchange.....	53
XIV	Disproportionation of Arylsulfinic Acids.....	71
XV	Disproportionation of Symmetrical Disulfides..	76

LIST OF FIGURES

Figure

- 1 Log of Rate of Disproportionation of
p-Substituted Benzenesulfinic Acids in
Acetic Acid-0.56 M Water-0.60 M
Sulfuric Acid at 80°C vs. σ^+ and σ 17
- 2 Plot of Log k versus 1/T for
p-Toluenesulfinic Acid in Acetic
Acid-0.60 M Sulfuric Acid-0.56 M
Water Solution..... 19
- 3 Reaction Coordinate-Free Energy Profile
Diagram. Electron-Withdrawing Substituent... 24
- 4 Reaction Coordinate-Free Energy Profile
Diagram. Electron-Donating Substituent..... 24
- 5 Reaction Vessel used in Kinetic Experiments.. 34
- 6 Plot of ln Q versus t(min.)..... 54

STUDIES OF THE MECHANISM OF THE
DISPROPORTIONATIONS OF SULFINIC ACIDS AND DISULFIDES

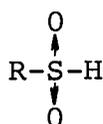
PART I: THE DISPORPORTIONATION OF SULFINIC ACIDS

INTRODUCTION

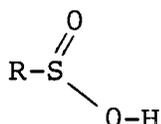
Sulfinic acids have been known since the late 1850's. Kolbe (50) and Kalle (41) both reported preparations for benzenesulfinic acid in 1860 and 1861 respectively. Otto (63) prepared p-toluenesulfinic acid in 1868.

The Structure of Sulfinic Acids

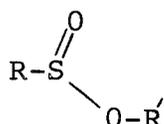
The structure of sulfinic acids has been a subject of much discussion for some time. Bredereck and co-workers (8) have used ultraviolet spectral data to show that sulfinic acids exist in the ester form (II) in both polar and non-polar solvents. Methyl and ethyl p-toluene



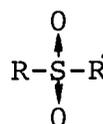
I



II



III



IV

sulfinates (III) absorb at 250 $\text{m}\mu$ ($\log \epsilon = 3.5$) whereas the alkyl aryl sulfones (IV) show multiple peaks in the 260-270 $\text{m}\mu$ region ($\log \epsilon = 2.6-2.7$).

Some Russian workers (34) felt that the infrared spectra showed an S-H, indicative of structure (I). Mercaptans are known to absorb in the 2500-2600 cm^{-1} range (4, p. 350) and a similar peak has been observed with sulfinic acids by many authors (8,21,34). Detoni and Hadži (21) report that this peak is not due to an S-H bond, but that it is due to very pronounced hydrogen bonding. On dilution these authors have observed a new,

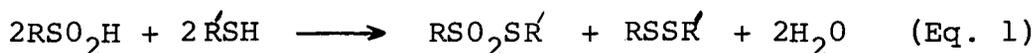
sharp peak at 3700 cm^{-1} which occurs at the expense of the lower peak and is probably due to the monomer. Also, the $2500\text{--}2600\text{ cm}^{-1}$ peak absorbs only very weakly in the Raman spectra (34) which indicates an O-H bond rather than an S-H bond, as S-H bonds are known to absorb strongly in the Raman spectra (15,21).

Chemical Reactions of Sulfinic Acids

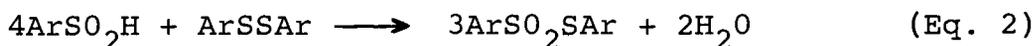
Both sulfinic acids and their salts can undergo air oxidation to the higher oxidation state sulfonic acids (76).

Sulfinic acids may also undergo auto-oxidation, probably via a radical reaction (39, p. 109). In addition, they can also undergo a variety of other oxidation reactions (30).

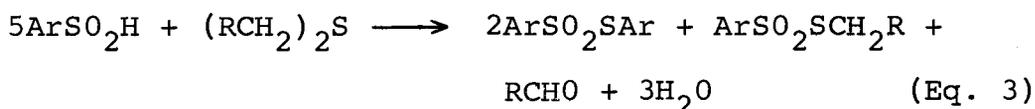
Von Braun (77) observed the following reaction of a sulfinic acid with a mercaptan:



Kice and Bowers (44) studied the reaction p-toluenesulfinic acid and p-tolyl disulfide. The stoichiometry of the reaction is:

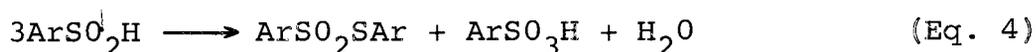


These authors (45) also discovered the reaction of p-toluenesulfinic acid with certain alkyl sulfides:



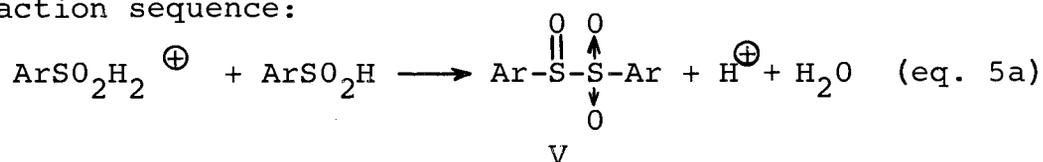
Disproportionation of Sulfinic Acids

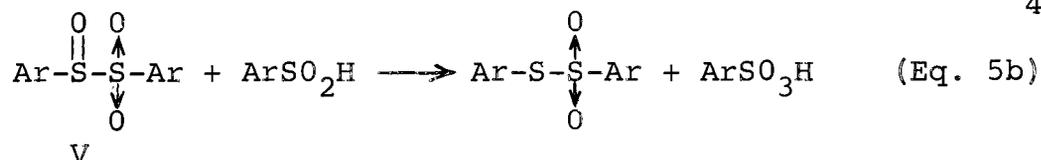
Otto and Pauley (68) were the first to study the disproportionation reaction of sulfinic acids. They found that on heating in solution p-toluenesulfinic acid underwent a disproportionation having the following stoichiometry:



They found that the reaction could be suppressed by the addition of water. For this reason they felt, incorrectly, as it later turned out, that the disproportionation was probably an equilibrium. Otto (64) also reported that benzenesulfinic acid and p-chlorobenzenesulfinic acid disproportionate according to Equation 4. Fromm and Seixas Palma (29) reported that α -toluenesulfinic acid disproportionated in a manner analogous to that reported for p-toluenesulfinic acid.

No serious attempt was made to elucidate the mechanism of the reaction until quite recently. In 1958 Brederick and co-workers (9) reported that the disproportionation of p-toluenesulfinic acid was acid catalyzed in dioxane solution. They assumed that the starting material was partially protonated and proposed the following reaction sequence:



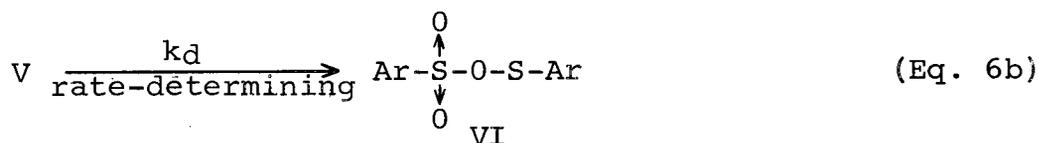
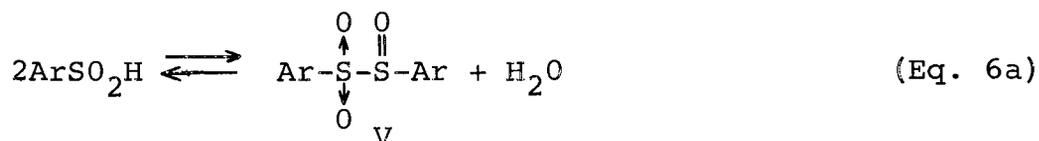


As additional support for this mechanism they showed that in anhydrous dioxane p-toluenesulfinic acid and p-toluenesulfinyl p-tolyl sulfone (V) do react to give at least 75% of the amount of thiol-sulfonate (ArSO_2SAr) expected from the stoichiometry of Equation 5b.

More recently, Kice and Bowers (43) have studied the kinetics of the disproportionation of p-toluenesulfinic acid in acetic acid-sulfuric acid-water solution. They found the reaction to be clearly second order in sulfinic acid. They also found that added strong acid (H_2SO_4) had only a modest accelerating effect on the rate. On the other hand, an increase in the stoichiometric concentration of water in acetic acid-water solvent led to a very large decrease in the rate of disproportionation. In addition, some studies done with p-toluenesulfinyl p-tolyl sulfone showed that a direct reaction between V and sulfinic acid (such as Equation 5b) cannot be involved in the disproportionation reaction in acetic acid solution. Specifically, Kice and Bowers showed that in the acetic acid-sulfuric acid-water solvent system the sulfinyl sulfone hydrolyzes very rapidly with formation of two molecules of sulfinic acid. Any reaction of added sulfinic acid with the sulfinyl sulfone was shown to be

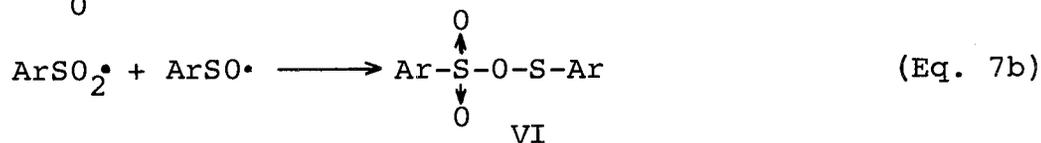
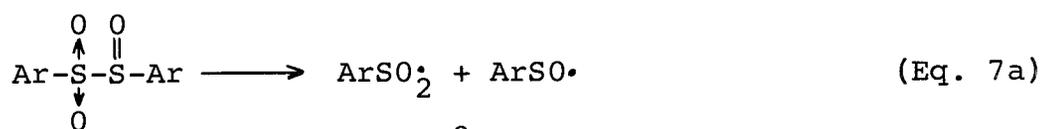
slow compared to the rate of hydrolysis of V. Therefore, were Brederick's mechanism to apply, the reaction of the sulfinyl sulfone with the sulfinic acid would have to be the rate-determining step; but in such a situation the overall kinetics of the reaction would then be third order with respect to sulfinic acid, rather than second order as has been found.

Although Kice and Bowers showed that the disproportionation does not involve a direct reaction between sulfinyl sulfone and sulfinic acid, they did not rule out the sulfinyl sulfone as the possible key intermediate. In fact, they proposed the following mechanism involving rate-determining unimolecular decomposition of V for the disproportionation of p-toluenesulfinic acid in acetic acid-sulfuric acid-water solution:



Strong support for the correctness of this mechanism was provided by a subsequent investigation by Kice and Pawlowski (46,47) of the thermal decomposition of the sulfinyl sulfone (V). They showed that in anhydrous

dioxane V undergoes a quite rapid unimolecular decomposition whose rate is the same in the presence of sulfinic acid as in its absence. Although added sulfinic acid does not influence the rate, it does alter the decomposition products, and in its presence the overall stoichiometry of the decomposition of V is given by Equation 5b. These results are what one would expect for a mechanism in which sulfinic acid is consumed, not by a direct reaction with V, but rather by a fast reaction with an intermediate or intermediates formed from the rate-determining decomposition of the sulfinyl sulfone. Other studies by Kice and Pawlowski suggested that the initial step in the decomposition of V was a homolytic cleavage of the S-S bond in the sulfinyl sulfone. They accordingly proposed the following mechanism for the sulfinyl sulfone decomposition:



The sulfenyl sulfonate (VI) seems to be a very reasonable intermediate as its reactions should be analogous to the reactions of sulfenyl halides which are known to react very rapidly with sulfinic acids to form similar products.

In the absence of added sulfinic acid the decomposition of V still yields about one mole of ArSO_3H per mole of V but the amount of thiolsulfonate produced is much lower and a sizeable amount of p-tolyl disulfide is produced. How the disulfide and sulfonic acid are formed under these conditions is not clear at the present time.

Thermodynamic activation quantities were determined (47) for the sulfinyl sulfone decomposition in both dioxane and acetonitrile and are shown in Table I. Though these quantities change quite markedly, the change

Table I

Thermodynamic Quantities for Sulfinyl
Sulfone Decomposition *

Solvent	ΔH^\ddagger	ΔS^\ddagger
Dioxane	27.6 kcal	+11.2 e.u.
Acetonitrile	24.8 kcal	+ 1.8 e.u.

* Data from Reference 47.

in one almost completely compensates for the change in the other so that the rate of decomposition of V is not appreciably solvent dependent.

Kice and Pawlowski also studied the effect of the structure of the aryl group on the rate of decomposition of V. These results are shown in Table II. These results indicate that electron withdrawing substituents tend to accelerate the reaction slightly, whereas

electron donating substituents hinder the reaction.

Table II

Effect of Structure on the Rate of Sulfinyl Sulfone Decomposition *

Ar	$k_d \times 10^4 \text{ sec}^{-1} **$
p-CH ₃ C ₆ H ₄ -	3.9
C ₆ H ₅ -	4.9
p-ClC ₆ H ₄ -	6.4
β-Naphthyl-	6.4

* Data from Reference 47.

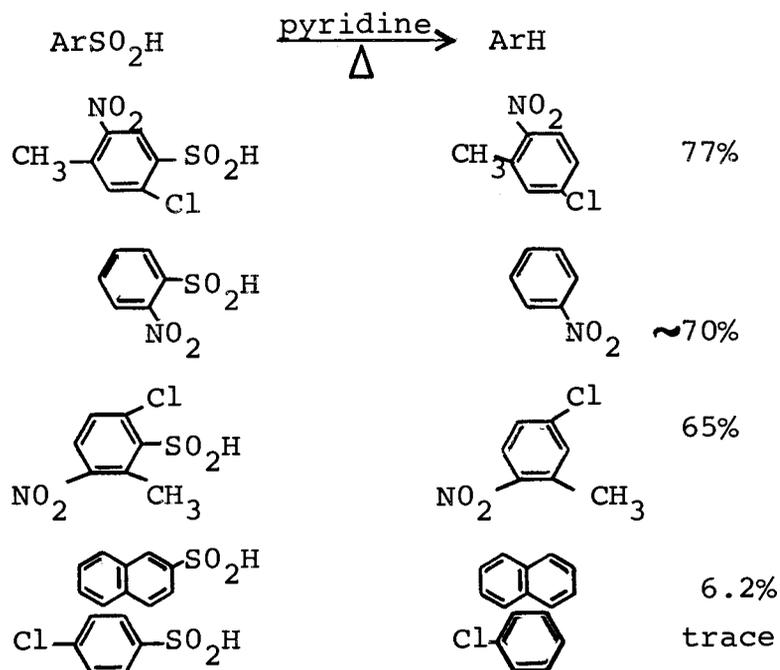
** All determined at 49.3° C.

Dann and Davies (19) showed that sulfinic acids with electron withdrawing groups were so stable, in some cases, that they lost sulfur dioxide on heating in solution rather than undergoing the normal disproportionation reaction. Some of their results are shown in Table III.

These results seem to indicate that similar substituents have opposite effects on the stability of sulfinyl sulfones as compared to the stability of the analogous sulfinic acids in the normal disproportionation reaction. It was therefore of interest to determine quantitatively the effect of structure on the rate of disproportionation.

Table III

Ease of SO₂ Loss of Some Sulfinic Acids
on Heating in Solution *



* Data from Reference 19.

Hampton (49) has done some preliminary work concerning the effect of structure on the rate of disproportionation. His results are shown in Table IV.

Table IV

Effect of Structure on Sulfinic Acid
Disproportionation *

Ar	$k_2 \times 10^{-3} \text{ mole}^{-1} \text{ sec}^{-1}$
C ₆ H ₅ -	1.91
p-ClC ₆ H ₄ -	1.05

* Data from Reference 49;
0.56 M H₂O, 0.60M H₂SO₄, Temp. = 80.3°C.

Hampton also did work to show that these differences in rate were truly due to structural effects rather than to differences with substituent in the effect of water or sulfuric acid concentration on rate. The results of the studies on the effects of varying sulfuric acid concentration and water concentration are shown in Tables V and VI.

Table V

Effect of Sulfuric Acid Concentration on Rate of Disproportionation *

Ar	H ₂ SO ₄ concentration			
	0.0 M	0.3 M	0.6 M	1.2 M
	k ₂ / (k ₂ at 0.6M H ₂ SO ₄) for ArSO ₂ H			
C ₆ H ₅ -	0.34	0.72	(1.0)	1.52
p-ClC ₆ H ₄ -	0.38	0.76	(1.0)	1.55
p-CH ₃ C ₆ H ₄ -**	0.37	0.73	(1.0)	1.61

* Data from Reference 49; 0.56 M H₂O in acetic acid.

** Data from Reference 43; 0.56 M H₂O in acetic acid.

Table VI

Effect of Water on Rate of
Disproportionation *

	H ₂ O concentration		
	0.56 M	1.12 M	1.68 M
Ar	$k_1/k_2 / (k_2 \text{ at } 0.56\text{M H}_2\text{O}) \text{ for ArSO}_2\text{H}$		
C ₆ H ₅ -	(1.0)	0.43	0.30
p-ClC ₆ H ₄ -	(1.0)	0.45	0.28
p-CH ₃ C ₆ H ₄ - **	(1.0)	0.39	0.22

* Data from Reference 49; C₆H₅SO₂H and p-ClC₆H₄SO₂H in acetic acid-0.60 M H₂SO₄.

** Data from Reference 43; CH₃C₆H₄SO₂H in acetic acid - 1.0 M H₂SO₄.

Since the relative rates for a particular set of conditions are in good agreement, the changes in rate in Table IV are almost certainly due to structural rather than solvent effects.

Hampton's results are further evidence that the effect of structure on the rate of disproportionation is just the opposite of what was observed in the study of the sulfinyl sulfones. Therefore it was felt that a more thorough study of the effect of structure would be of value.

Besides such studies a determination of the thermodynamic activation parameters for the overall reaction was thought to be of value. Knowing these quantities

one should be able to predict the energy requirements for the reaction profile. Similarly, this information can be used to show how the sulfinyl sulfone may be related to the sulfinic acid in a thermodynamic sense.

RESULTS

Kinetics of Disproportionation

The disproportionation of several aromatic sulfinic acids has been studied in acetic acid-sulfuric acid-water solutions. The reaction was followed by removing aliquots of the solution from the reaction vessel, quenching with water, then titrating the remaining sulfinic acid with a standard solution of sodium nitrite. A different method had to be devised for following the reaction of the p-nitro compound. This alternate procedure involved the titration of the aliquot from the reaction with standard ceric sulfate using a potentiometric end point. The thiolsulfonate from the disproportionation is not oxidized to any appreciable amount under these conditions; therefore the determination of sulfinic acid is quantitative. This method is described in greater detail in the Experimental section.

In all cases the disproportionation reaction followed good second-order kinetics. The results of the various runs are shown in Table VII. The results are reported in terms of k_2 which is the experimentally measured second-order rate constant for the reaction.

Table VII

Kinetics of the Disproportionation of
Aromatic Sulfinic Acids in Acetic Acid

Ar ,	M	Temp., °C	(H ₂ O) , *	M (H ₂ SO ₄) , *	M k ₂ × 10 ³ , M ⁻¹ sec ⁻¹
p-BrC ₆ H ₄ ,	0.10	80.3	0.56	0.60	1.01
p-CH ₃ C ₆ H ₄ ,	0.10	59.6	0.56	0.60	0.43
		69.8			1.45
		80.3			4.50
		89.4			12.0
		69.8	0.56	0.30	1.03
		80.3			3.20
p-CH ₃ OC ₆ H ₄ ,	0.10	80.3	0.56	0.60	10.7
p-O ₂ NC ₆ H ₄ ,	0.10	80.3	0.56	0.60	0.36
β-Naphthyl,	0.10	69.8	0.56	0.60	2.20
		80.3			6.95

* Stoichiometric concentrations used throughout.

Effect of Structure on Rate of Disproportionation

Hampton (49), in his preliminary work, has shown that differences in rate with substituent are due to structural effects rather than to differences with substituent in the influence of water or sulfuric acid concentration on the rate. The results of this preliminary work are shown in Tables IV and V. In the present study the rates of disproportionation of various substituted aromatic sulfinic acids were determined in acetic acid containing 0.60 M sulfuric acid and 0.56 M water at 80°C. Under these conditions the rates were conveniently measurable for all of the compounds studied. The

effects of electron donating and electron withdrawing substituents were studied. Some work was done on the effect of bulk of the aromatic group. Table VIII shows the relative rates of disproportionation of the various sulfinic acids studied, benzenesulfinic acid being taken as the reference compound.

Table VIII

Effect of Structure on Rate of Disproportionation of Aromatic Sulfinic Acids *

Ar	$k_2\text{ArSO}_2\text{H}/k_2\text{PhSO}_2\text{H}$ **
p-CH ₃ OC ₆ H ₄ -	5.4
p-CH ₃ C ₆ H ₄ -	2.36
C ₆ H ₅ -	(1.0)
p-ClC ₆ H ₄ -	0.55
p-BrC ₆ H ₄ -	0.53
p-O ₂ NC ₆ H ₄ -	0.19
β-Naphthyl-	3.64

* Data for C₆H₅- and p-ClC₆H₄- from Hampton (49).

** All data at 80°C in acetic acid-0.60M sulfuric acid-0.56M water.

It is obvious that with para-substituted benzenesulfinic acids electron-donating groups accelerate the disproportionation reaction and electron-withdrawing groups retard the reaction.

The Hammett equation (38, p. 86) is an extremely useful relationship for assessing the effect of changes in structure on reactivity in aromatic systems. In this

equation (Equation 8) k and k_0 refer to the reaction rate constants for the substituted and unsubstituted compounds, respectively; σ is a substituent constant, which is a measure of the electron-donating or electron-withdrawing power of the substituent; and ρ , known as the reaction constant, is a measure of the sensitivity of the reaction to changes in electron density at the reaction site induced by the substituent. Hammett originally felt that such a linear free energy relationship as Equation 8 would only be observed in cases where there was not

$$\log \frac{k}{k_0} = \rho\sigma \quad (\text{Eq. 8})$$

important resonance interaction between substituent and reaction site in either reactants or transition state. However, further investigations (11,75) have shown that many cases where there can be direct resonance interaction between the substituent group and the reaction center also give good correlation with Equation 8 provided that different sets of σ values are used. For cases where the resonance interaction formally involves electron-donation from substituent to reaction site σ^+ values are used. For cases where the interaction involves electron-donation from reaction site to substituent σ^- values are appropriate.

Figure 1 shows a plot of $\log k_2$ versus σ and σ^+ for the disproportionation of the various benzenesulfinic

acids studied. It is evident that while the rates are

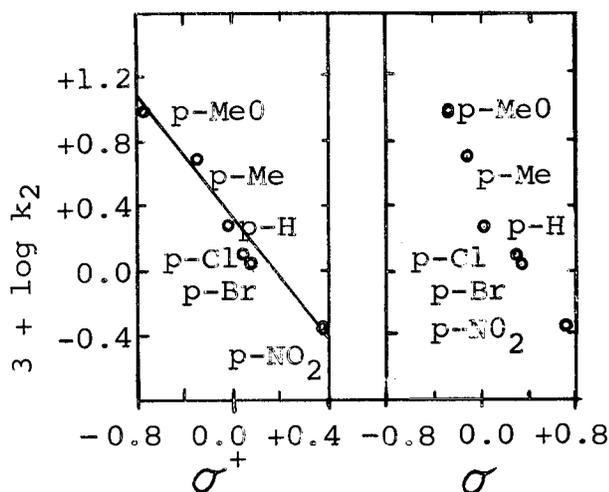


Figure 1. Log of rate of disproportionation of p-substituted benzenesulfonic acids in acetic acid-0.56 M water-0.60 M sulfuric acid at 80°C vs. σ^+ and σ .

correlated fairly well by σ^+ ($\rho = -1.1$), there is no linear correlation between $\log k$ and σ . The significance of this result will be taken up in the discussion.

Determination of Activation Parameters

Since Kice and Pawlowski (47) determined the activation parameters for p-toluenesulfinyl p-tolyl sulfone in two solvents, it was felt that the corresponding quantities for p-toluenesulfonic acid should be determined. Consequently, studies of the rate of disproportionation of this compound over a forty degree temperature range were done. This work was done in acetic acid-0.60 M sulfuric acid-0.56 M water solution and the data is shown in Table VII. A plot of $\log k$ versus $1/T$

(Figure 2) shows a linear relationship. Using this data, ΔH^\ddagger is calculated to be 26.4 kcal (30, p. 23-24).

Experiments at 70°C and 80°C suggest that essentially the same linear relationship is shown in acetic acid-0.30 M sulfuric acid-0.56 M water solution. The enthalpy of activation apparently does not depend on the acidity of the solution to any great extent. Work with the β -naphthyl compound at these two temperatures indicated that the enthalpy of activation for β -naphthalenesulfinic acid was nearly the same as that obtained for p-toluenesulfinic acid. The ratios of k_2 at 70°C to k_2 at 80°C are shown in Table IX.

Table IX

Ratios of k_2 at 80°C to k_2 at 70°C for
Varying Sulfuric Acid Concentration and
Aryl Group

Ar	$[H_2SO_4]$	$[H_2O]$	k_2 80°C/ k_2 70°C
p-CH ₃ C ₆ H ₄	0.60	0.56	3.10
	0.30	0.56	3.11
β -Naphthyl	0.60	0.56	3.11

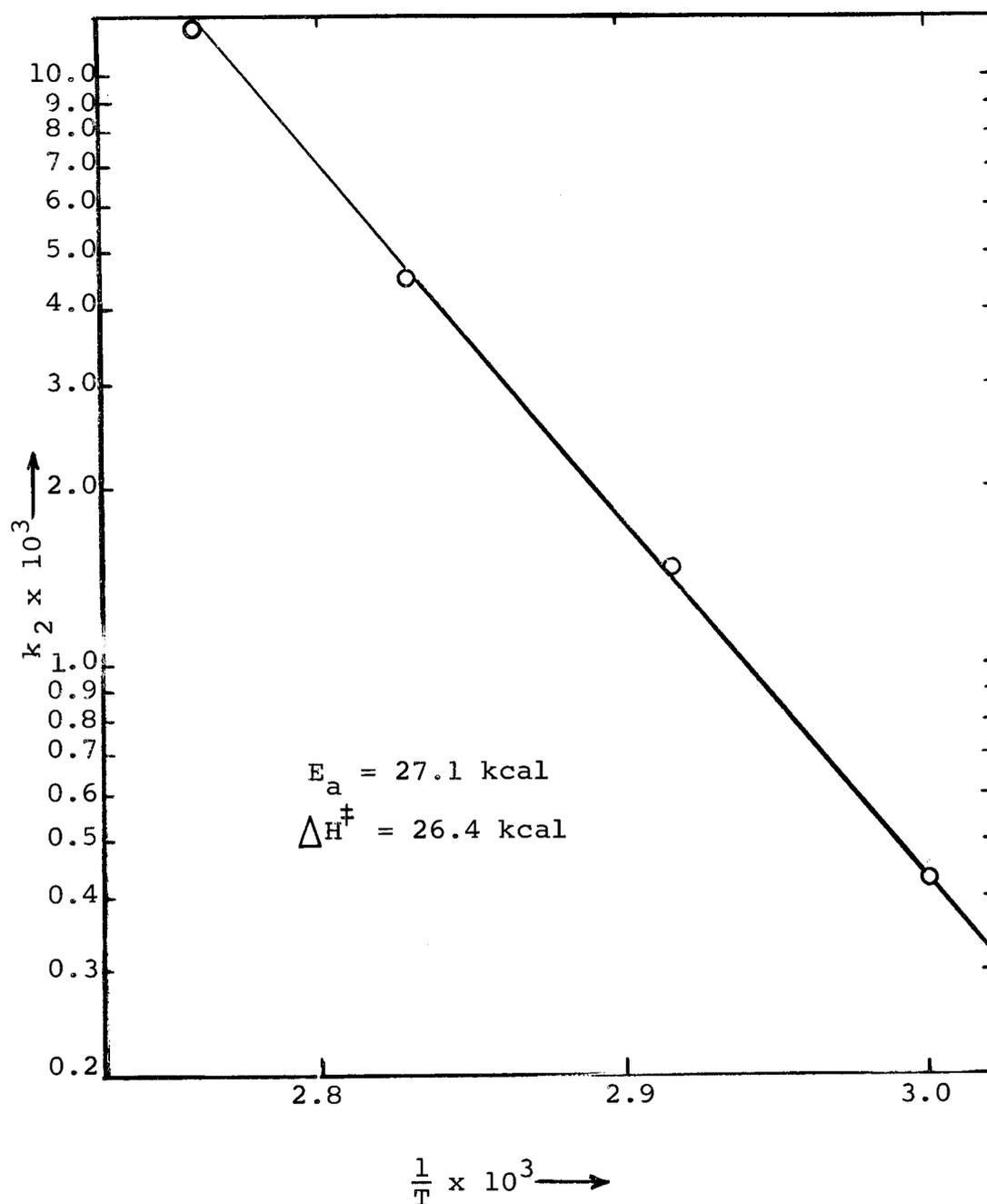


Figure 2. Plot of $\log k$ versus $1/T$ for p-toluenesulfinic acid in acetic acid-0.60 M sulfuric acid-0.56 M water solution.

DISCUSSION

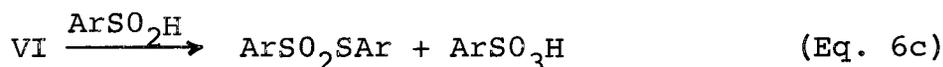
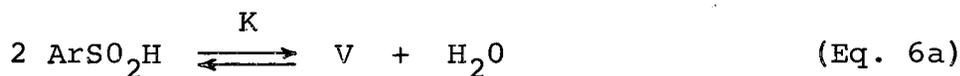
Thermodynamic Activation Parameters

The data given in Table VII for the disproportionation of p-toluenesulfinic acid at several temperatures allow one to calculate ΔH^\ddagger for the reaction. These calculations yield a ΔH^\ddagger of 26.4 kcal. for the acetic acid-0.60 M sulfuric acid-0.56 M water solutions. Data for the decomposition of p-toluenesulfinic acid at two temperatures in acetic acid-0.30 M sulfuric acid-0.56 M water solution indicates that ΔH^\ddagger does not depend on the acidity of the solution. Experiments with 2-naphthalenesulfinic acid show that ΔH^\ddagger does not change to any great extent with variation of the structure of the aryl group.

Since the ΔH^\ddagger for the sulfinyl sulfone decomposition (47) showed a considerable variation with solvent (dioxane, ΔH^\ddagger was 27.6 kcal.; acetonitrile, ΔH^\ddagger was 24.8 kcal. Table I), we cannot say positively what the ΔH^\ddagger is for the decomposition of the sulfinyl sulfone in acetic acid solution. It is fairly certain that almost all of the ΔH^\ddagger for the sulfinic acid disproportionation must be due to the ΔH^\ddagger of the sulfinyl sulfone decomposition. If this is the case, then the enthalpy change which one attributes to the first equilibrium must be quite small. This would indicate that the first

equilibrium is not as unfavorable as one might predict unless there is an unusually large negative entropy for the first reaction. Although the concentration of sulfinyl sulfone is never large enough to cause a deviation from second-order kinetics, it may, in some cases, amount to as much as a few percent of the total sulfinic acid concentration present.

One can use a somewhat different approach to make semi-quantitative calculations to show what amount of the sulfinyl sulfone may be present in the solutions which were studied. When one considers the mechanism shown in Equation 6 a-c it is apparent that one can express the



concentration of the sulfinyl sulfone (V) in terms of the equilibrium constant, K, the sulfinic acid concentration, $[\text{ArSO}_2\text{H}]$, the activity coefficients, f_i , of the compounds, and the activity of water, $a_{\text{H}_2\text{O}}$. When one

$$[\text{sulfinyl sulfone}]_{\text{equilibrium}} = \frac{K[\text{ArSO}_2\text{H}]^2 f_{\text{ArSO}_2\text{H}}^2}{f_{\text{V}} a_{\text{H}_2\text{O}}} \quad (\text{Eq. 9})$$

assumes a steady state approximation for the intermediate

sulfinyl sulfone one can derive the following expression for the rate of disproportionation. The experimentally

$$-\frac{d[\text{ArSO}_2\text{H}]}{dt} = \frac{3k_d K [\text{ArSO}_2\text{H}]^2 f_{\text{ArSO}_2\text{H}}^2}{f_* 6b a_{\text{H}_2\text{O}}} \quad (\text{Eq. 10})$$

measured rate constant, k_2 , is then given by Equation 11.

$$k_2 = \frac{3k_d K f_{\text{ArSO}_2\text{H}}^2}{f_* 6b a_{\text{H}_2\text{O}}} \quad (\text{Eq. 11})$$

k_1 , the experimentally measured (47) first-order rate constant for the decomposition of the sulfinyl sulfone is given by

$$k_1 = \frac{k_d f_V}{f_* 6b} \quad (\text{Eq. 12})$$

Equation 13 gives the ratio of k_2 to k_1 as determined

$$\frac{k_2}{k_1} = \frac{3K f_{\text{ArSO}_2\text{H}}^2}{f_V a_{\text{H}_2\text{O}}} \quad (\text{Eq. 13})$$

by use of Equations 11 and 12. Substitution of this expression into Equation 9 gives the following relationship between the equilibrium concentration of sulfinyl sulfone and the ratio (k_2/k_1) (Equation 14).

$$[\text{sulfinyl sulfone}]_{\text{equil.}} = \frac{k_2}{3k_1} [\text{ArSO}_2\text{H}]^2 \quad (\text{Eq. 14})$$

Kice and Pawlowski (47) have shown that although ΔH^\ddagger for the sulfinyl sulfone decomposition is quite solvent dependent, the rate constants for the decomposition of V do not vary to any great extent with solvent. It was

therefore assumed that one could use the rate constant for the sulfinyl sulfone decomposition in dioxane as a reasonable k_1 value for the decomposition of V in acetic acid. From the rate data for the p-toluenesulfinyl p-tolyl sulfone decomposition in dioxane (47) one can estimate k_1 at 80°C to be about $2.2 \times 10^{-2} \text{ sec}^{-1}$ k_2 for p-toluenesulfinic acid in acetic acid-0.60 M sulfuric acid-0.56 M water at 80°C is $4.3 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$ (Table VII). When these data are used in Equation 14 one calculates the equilibrium concentration of sulfinyl sulfone in a solution which is 0.10 M with respect to sulfinic acid to be about $6.5 \times 10^{-4} \text{ M}$. Similar calculations can be carried out with data for the p-chloro sulfinyl sulfone yielding a k_1 at 80°C for that compound of about $3.7 \times 10^{-2} \text{ sec}^{-1}$ k_2 is $1.05 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$ at this temperature (Table VII). These data indicate that the p-chlorobenzenesulfinyl p-chlorobenzene sulfone reaches an equilibrium concentration of about $9.5 \times 10^{-5} \text{ M}$. These approximate calculations show that the compound with the more electron donating substituent is the one which has the higher concentration of sulfinyl sulfone at equilibrium.

This data for the sulfinic acid disproportionation combined with the data from the work of Kice and Pawlowski (47) on the sulfinyl sulfones allows one to

graph, at least qualitatively, a reaction coordinate-free energy profile diagram for the reaction. Figures 4 and 5 show the graphs that one may envision for the cases of electron-withdrawing and electron-donating substituents, respectively.

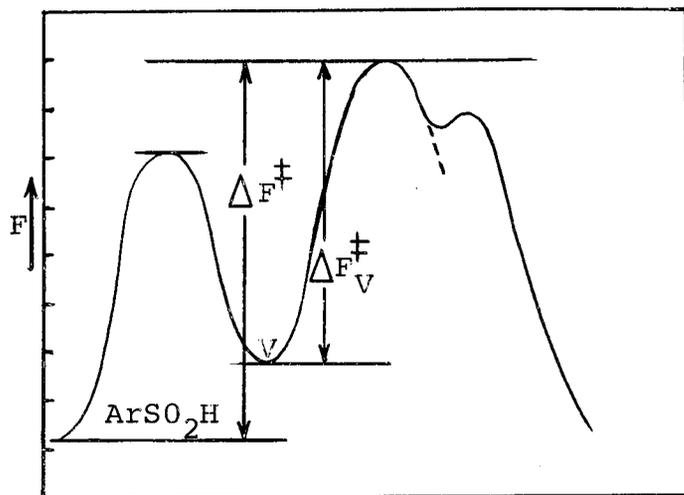


Figure 3. Reaction Coordinate \longrightarrow
Electron-Withdrawing Substituent

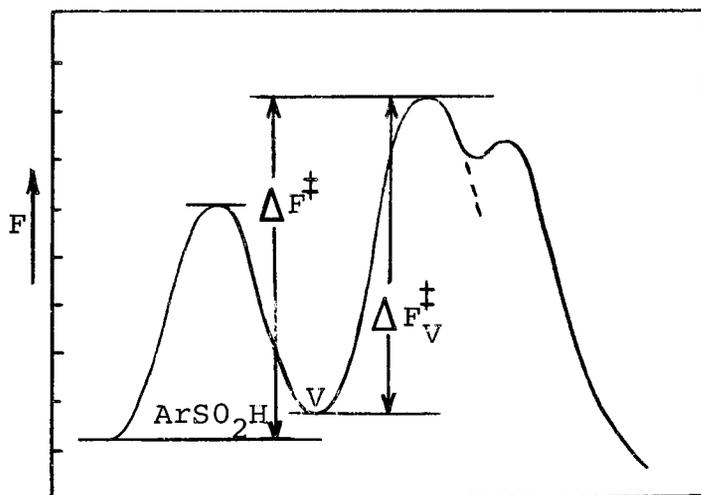
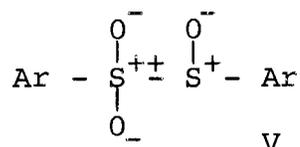


Figure 4. Reaction Coordinate \longrightarrow
Electron-Donating Substituent

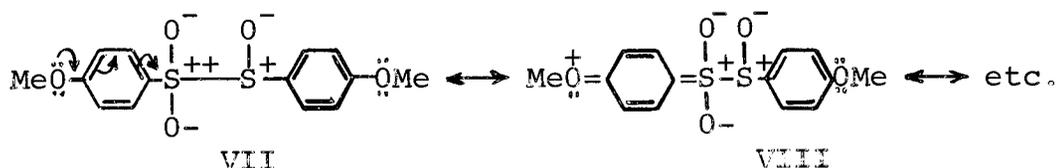
Effect of Structure on Sulfinic Acid Disproportionation

It has been shown that electron withdrawing groups retard and electron donating groups accelerate the disproportionation of para-substituted benzenesulfinic acids. This is just the opposite of what was found by Kice and Pawlowski (47) for the decomposition of the proposed intermediate in the reaction, the sulfinyl sulfone. To explain these differences one must consider the mechanism of the reaction as proposed by Kice and Bowers (43) (Equation 6 a-c). The formation of the sulfinyl sulfone (V) should be favored by electron donating substituents



which will decrease the repulsion between the electron deficient sulfur atoms. In the cases where the formation of the sulfinyl sulfone is favored, the rate of reaction will be faster due to the direct dependence of the rate on the concentration of the intermediate sulfinyl sulfone. In the sulfinyl sulfone studies (47), the compounds with electron withdrawing substituents decomposed slightly faster due to their inherent instability, though there was less than a factor of two in the difference in rate between the p-chloro and the p-tolyl compounds. It may be stated that most of the differences in rate with substituent for the disproportionation of para-substituted

benzenesulfinic acids are due to differences in the position of the initial equilibrium as caused by stabilizing or destabilizing effects of the substituent. The correlation of $\log k$ with σ^+ substituent constants is a further indication of this. One may show direct resonance stabilization of the sulfinyl sulfone when the substituent is an electron-donating one such as a para-methoxy group which can have resonance forms such as VIII that will



make important contributions to the stability of the intermediate. When the product of the first equilibrium can be stabilized in this manner, the concentration of sulfinyl sulfone present will be significantly larger than when this type of stabilization does not exist. The rate-determining step in the disproportionation reaction is the unimolecular decomposition of the intermediate sulfinyl sulfone. This decomposition reaction is relatively insensitive to the substituent on the aryl group which indicates that both the sulfinyl sulfone and the transition state for the sulfinyl sulfone decomposition are stabilized to about the same extent by electron donating groups. For these reasons the overall rate of disproportionation depends almost entirely on the

position of the first equilibrium. It may be concluded that the compounds with electron donating substituents disproportionate more rapidly due to the favorable position of the first equilibrium. This is in agreement with the conclusions one might surmise from the semi-quantitative calculations which were shown in the preceding section.

EXPERIMENTAL

1.) Preparation of Sulfinic Acids

p-Bromobenzenesulfinic Acid. The sodium salt of the acid was prepared by reduction of p-bromobenzenesulfonyl chloride with sodium sulfite in basic solution. The reduction was carried out following the directions of Bäder and Hermann (3). The sodium salt was dissolved in water, filtered, and 6 N sulfuric acid was used to precipitate the sulfinic acid. The product was filtered off and recrystallized at least twice using ether and hexane. This product titrated 98% purity and melted at 112-114°C. Literature value (13) 113-114°C.

p-Methoxybenzenesulfinic Acid. Anisole was reacted with chlorosulfonic acid in chloroform at -8°C to prepare p-methoxybenzenesulfonyl chloride. The reaction was carried out according to the directions of Morgan and Cretcher (57) and gave a 45% yield. The sulfonyl chloride was reduced with sodium sulfite in basic solution to obtain the sodium salt in solution. This solution was treated with an acidic solution of ferric chloride which precipitated the ferric salt (67) of the acid. The ferric salt was isolated by gravity filtration and stored under refrigeration until just before use. The ferric salt was destroyed by treatment with cold concentrated sodium hydroxide solution, the ferric oxide filtered off

by gravity filtration, and the sulfinic acid was isolated by filtration after treatment of the chilled solution with cold 6 N sulfuric acid. The acid was then dried and recrystallized from ether by precipitation with hexane. Two recrystallizations were sufficient to give a product that titrated 95% purity and had a melting point of 95-97°C. Literature value (67) 98-99°C.

2-Naphthalenesulfinic Acid. The method of Bäder and Hermann (3) was used for the reduction of 2-naphthalene-sulfonyl chloride with sodium sulfite in basic solution to prepare the relatively insoluble salt of 2-naphthalenesulfinic acid. The sodium salt was dissolved in water, filtered, and 6 N sulfuric acid was used to precipitate the acid. The acid was filtered off and dissolved in ether with stirring to obtain a saturated solution. This solution was dried over anhydrous calcium sulfate and then the acid was precipitated by addition of five volumes of hexane to one volume of ether solution. Recrystallization of the first fraction which precipitated by the same method yielded compound which titrated 97-99% purity and melted at 103-104.5°C. Literature value (66) 105°C.

p-Nitrobenzenesulfinic Acid. The sodium salt was prepared by basic reduction of the sulfonyl chloride using sodium sulfite (3). The acid was prepared by

addition of cold concentrated hydrochloric acid to a concentrated solution of the salt in water. The product was filtered off and dissolved in ether, and the solution was dried over anhydrous calcium sulfate. Three volumes of hexane were added to one volume of the ether solution to precipitate the acid. A second recrystallization using the same procedure produced compound that titrated 99-100% purity and had a melting point of 157-158°C. Literature value (80) 159°C.

p-Toluenesulfinic Acid. The sodium salt was prepared (31, p. 492) by reduction of the sulfonyl chloride with zinc metal. The sodium salt was recrystallized from water and dried over phosphorus pentoxide under vacuum. The acid was prepared as needed from the salt by dissolving the salt in water, precipitating with cold 6 N sulfuric acid, and filtering off the product. This product was dried and dissolved in ether. The ether solution was filtered and one volume of hexane was added to one volume of the ether solution to precipitate the acid. The acid was recrystallized a second time by the same procedure and the product isolated titrated 97% purity and had a melting point of 84-85°C. Literature value (13) 84-85°C.

2.) Preparation of Stock Solutions

Glacial acetic acid (Baker and Adamson, reagent grade) was refluxed with 10% of its volume of acetic

anhydride for 24 hours to remove all traces of water. The acid was then slowly distilled through an Oldershaw column (20 theoretical plates) with only the large center fraction being used for kinetic experiments. The reflux ratio was kept at 15:1 for the distillation.

A solution of 0.56 M water in acetic acid was prepared by weighing the necessary amount (10.091 grams) of water into a one-liter flask and adding purified acetic acid with shaking. The solution was not made up to volume with the purified acetic acid until the water-acetic acid solution had been allowed to equilibrate in a water bath at 20°C for thirty minutes.

Reagent grade sulfuric acid (Baker and Adamson) was used to prepare 250 ml. of 3 M sulfuric acid-0.56 M water in acetic acid solution. The concentrated sulfuric acid was calculated to contain 3.75% water, therefore 76.437 grams of sulfuric acid and 0.7947 grams of acetic anhydride were weighed into the volumetric flask. All solutions were equilibrated at 20°C for 30 minutes in a water bath before filling to volume. In this manner the stock solution of sulfuric acid-0.56 M water in acetic acid was made up with five to ten times the sulfuric acid concentration desired for the various kinetic runs.

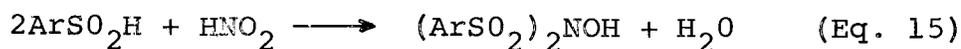
3.) Procedures for Following Kinetic Experiments with Substituted Arylsulfinic Acids

General Method for a Kinetic Run. An amount of arylsulfinic acid slightly in excess of that needed to prepare an 0.1 M solution was weighed out and dissolved in acetic acid-0.56 M water solution in a 50 ml. volumetric flask. The appropriate amount of 3 M sulfuric acid-0.56 M water in acetic acid solution was added to the volumetric flask and the solution was made up to the mark with the stock solution of 0.56 M water in acetic acid. All solutions were mixed thoroughly and all stock solutions were maintained at 20°C for at least 30 minutes prior to making up the solutions so as to maintain uniformity in volume. The solution was then transferred to the reaction flask and deaerated for five to ten minutes at room temperature with dry prepurified nitrogen gas. After deaeration the reaction vessel was placed in the constant temperature bath where it was allowed to equilibrate for five minutes before the first aliquot was removed at time zero. The aliquot (3 to 5 milliliters, depending on run) was quenched with one to three volumes of distilled water.

To remove an aliquot, nitrogen was bubbled rapidly into the solution and stopcock A (Figure 6) was turned 180° from the position shown in the diagram. After the

necessary volume of solution was forced up into the burette stopcock B was turned 45° in a clockwise direction and stopcock A was simultaneously returned to the position shown in the diagram. Thus a portion of the solution was trapped in the burette and the nitrogen pressure in the system was released. The flow of nitrogen through the solution was stopped and the sample was then removed from the burette by turning stopcock B another 45° in a clockwise direction. The remaining portion of the solution in the burette (about two milliliters, after sample removal) was then run out into a waste container and stopcock B was returned to its original position.

The sample, after being quenched with one to three times its volume of distilled water, was then titrated using an 0.2 N sodium nitrite solution. It is known (53) that two moles of sulfinic acid react with one mole of nitrous acid to form an N,N-di(arylsulfonyl) hydroxylamine in a quantitative manner (Equation 15). The tip of the burette containing the standard sodium nitrite



solution was immersed in the sample and the sample was stirred during titration by use of a magnetic stir bar and stirring motor. After each portion of nitrite was added to the solution a drop of the solution was removed

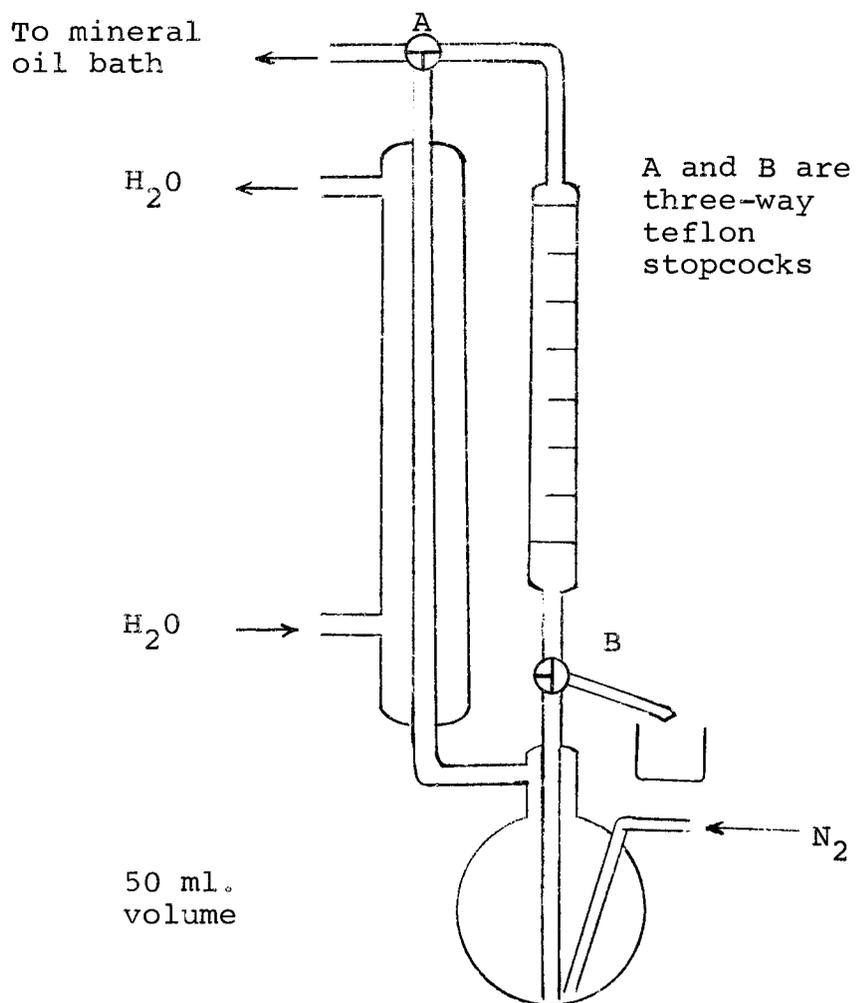


Figure 5. Reaction Vessel used in Kinetic Experiments

with a glass stirring rod and touched to a piece of starch-iodide paper placed on a white porcelain tile. A pinkish tinge, turning to purple, indicated that there was an excess of nitrite ion in the solution. This first faint color was taken as the endpoint.

Kinetic Studies of p-Methoxybenzenesulfinic Acid.

Because of the rapidity of the disproportionation, all of the samples were removed from the solution and quenched before they were titrated. In a trial run the normal procedure was followed and the rate constant was essentially the same as was found in the method described above.

Kinetic Procedure for p-Nitrobenzenesulfinic Acid.

Since the nitrite titration method does not work for p-nitrobenzenesulfinic acid, a different method for following the rate of disproportionation had to be devised. It was decided that a potentiometric method such as was described by Gringras and Sjöstedt (33) might be of use. A Beckmann Model 72 pH meter equipped with a combined platinum-calomel electrode was used for the determinations. A ceric sulfate solution was used as the oxidizing agent. Titration of 98% pure (determined by a nitrite titration) p-toluenesulfinic acid with the 0.05 N ceric sulfate solution showed that the method was quantitative. The potentiometric endpoint was found to be the

same as the endpoint determined with Ferroin indicator. The indicator could not be used in the kinetic runs because the large quantities of thiol-sulfonate which were present in the reaction mixture absorbed the indicator to such an extent that the color change at the endpoint was not sharp enough to be of use. The reaction was followed only to about 50% completion because the data determined at greater percent reaction completion were not accurate enough to be of value. The endpoints were determined by plotting ml. of ceric sulfate solution versus millivolts as indicated on the pH meter scale and taking the inflection point of the curve as the endpoint.

STUDIES OF THE MECHANISM OF THE
DISPROPORTIONATIONS OF SULFINIC ACIDS AND DISULFIDES

PART II: THE DISPROPORTIONATION OF DISULFIDES

INTRODUCTION

Disulfides have been known and used for some time. Zeise (79) first distilled ethyl disulfide from a mixture of potassium ethyl sulfate and barium disulfide in 1834. Other authors (14,58,61) also prepared ethyl disulfide by use of metal penta- and tri-sulfides. Calhoun (14) prepared methyl disulfide in 1846 and i-amyl disulfide was prepared by Muspratt (61) in 1851. Alkyl disulfides correspond in structure to alkyl peroxides. The disulfides are unlike the peroxides in that they are quite stable and easily prepared whereas the peroxides are unstable and even explosive in some cases.

Methods of Preparation

The most common method of preparation of a symmetrical disulfide involves the oxidation of the corresponding mercaptan. Common oxidizing agents include hydrogen peroxide, the halogens, and potassium ferricyanide. The lower molecular weight disulfides may be prepared by reaction of an alkyl halide with sodium disulfide followed by fractionation to separate the sulfide and the disulfide from the mixture of polysulfides which was formed. Disulfides may also be prepared by heating various organic compounds with sulfur. Mono-sulfides, tri-sulfides, and other products are also formed in the mixture, the proportions depending on the starting materials and on the conditions of heating. Disulfides

have been obtained from phenol (51), naphthols (62), saturated (55) and unsaturated (1) hydrocarbons, and benzene (32) by heating with elemental sulfur. Certain aldehydes and ketones are known to give good yields of disulfide when heated with hydrogen sulfide (15). Disulfides may be produced by treatment of a Grignard reagent with an excess of sulfur (74) or with sulfur monochloride (40). Sulfinic acids may be reduced to the disulfide under proper conditions (27,28). Sulfenyl chlorides, when treated with mercury, are known to produce disulfides (10).

One way in which unsymmetrical disulfides may be prepared is by treatment of a mixture of two mercaptans with an oxidizing agent such as bromine (65). Unfortunately this method gives a mixture of all three possible disulfides and requires the separation of the unsymmetrical disulfide from the two symmetrical compounds (54). Another method of preparation which succeeds in some cases is the reaction of a mercaptide with a Bunte salt to produce the unsymmetrical disulfide (56). Unfortunately, because of rapid equilibration reactions involving the mercaptide and the unsymmetrical disulfide first produced, this method also often gives a mixture of all three possible disulfides. The best and most common method of preparation of pure unsymmetrical disulfides is the reaction of a sulfenyl halide with a mercaptan (23). With this procedure one normally encounters no

problems with equilibration.

Reactions of Disulfides

Disulfides are quite easily reduced to mercaptans. There are many methods of reducing disulfides. Only a few will be mentioned here. A more complete list is given by Reid (69, p. 372-374). Disulfides are reduced to mercaptans by sodium hydrosulfide, sodium sulfide, sodium disulfide, or sodium polysulfide (7). Triphenyl phosphine, in the presence of water, reduces phenyl disulfide to the thiophenol (73). Disulfides can be reduced at the dropping mercury electrode in a polarographic analysis (35,78). Lithium aluminum hydride has been used for reducing disulfides (2) and sodium in ether has proved useful in the preparation of some alkyl mercaptides from alkyl disulfides (60).

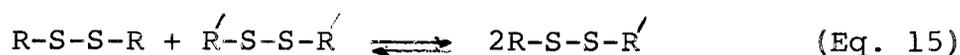
Theoretically, sulfenic acids, sulfinic acids, sulfonic acids, thiolsulfonates and disulfones may all be prepared by oxidation of disulfides using the correct conditions. Oxidation by hydrogen peroxide has been used to prepare products which are assumed to be formed from sulfenic acids (72). This method has also been used to prepare thiolsulfonates (71) and sulfinic acids (52). Hot nitric acid is used to prepare the sulfonic acid (20) and oxidation with permanganate has been used to prepare both the sulfonic acid and the α -disulfone (17).

Disulfides are known to undergo a disproportionation

reaction with mercaptan in alkaline alcoholic solution to yield another mercaptan and all possible disulfides. Gorin and co-workers (18) have recently done studies determining equilibrium constants for this disproportionation while other workers (6,54) have used the reaction for the preparation of various unsymmetrical disulfides.

Disulfide Disproportionation or Interchange Reaction

A mixture of two symmetrical disulfides can undergo the following equilibration:

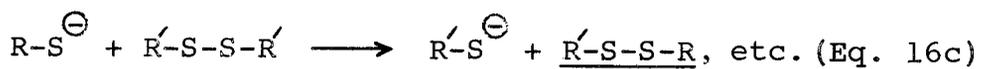
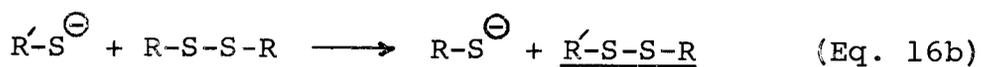
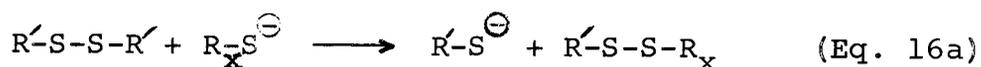


This disproportionation reaction is quite important and takes place fairly readily under a variety of conditions.

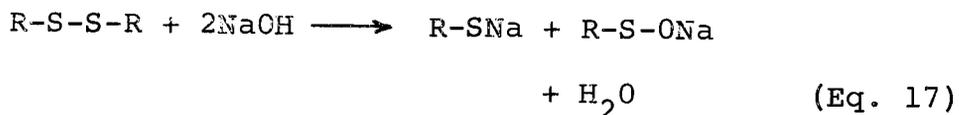
Birch and co-workers (6) found that methyl ethyl disulfide could not be distilled in the light because of disproportionation into methyl disulfide and ethyl disulfide, but distillation without equilibration went smoothly when they used a wrapped column. They also found that the compound decomposed from 99.2 mole percent purity to less than 90 mole percent purity in 20 hours in a clear bottle at room temperature, but could be stored for long periods of time in the absence of light. They proposed that the unsymmetrical disulfide disproportionation reaction goes by a chain mechanism which can be initiated by ultraviolet light. Disulfides are known to be very effective initiators of polymerization reactions

(6,25), usually under photolytic conditions.

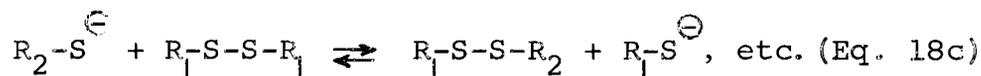
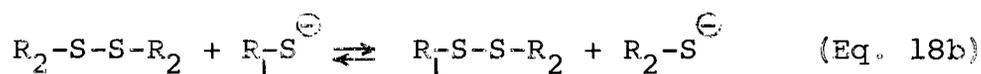
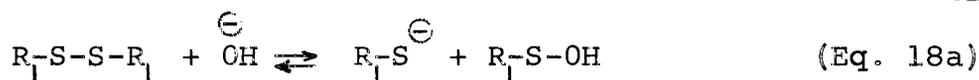
Much more common conditions for the disproportionation reaction are alkaline alcohol solutions containing a trace of mercaptan as a catalyst. These have been used in the preparation of unsymmetrical disulfides (6,54,70). Calvin (16) has proposed the following mechanism for the basic disproportionation of disulfides. He feels that a trace of mercaptan as catalyst might be produced by trace reduction or hydrolysis. It is known that disulfides are



split by alkali, and it is often stated that the first reaction is hydrolysis (24) (Equation 17). Ryle and



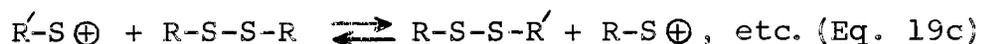
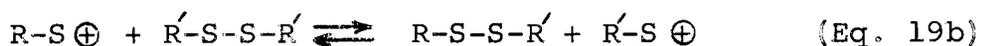
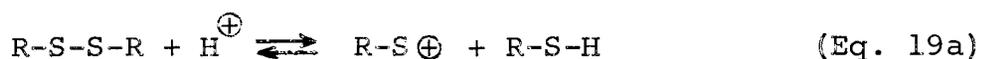
Sanger (70) proposed a mechanism which was almost identical to Calvin's. They found that when the disulfides were added to the basic solution there was a noticeable induction period before production of the unsymmetrical disulfide began, but when the disulfides were made up separately in basic solutions and allowed to stand for some time before mixing, there was no noticeable lag in the rate of formation of the unsymmetrical disulfide.



This indicated that the first equilibrium was established only rather slowly. These authors observed acceleration of the reaction rate with added thiol and inhibition of the reaction rate with addition of N-ethylmaleimide, a reagent which is known to react rapidly and quantitatively with thiols (26).

Ryle and Sanger (70) in 1954 and Benesch and Benesch (5) in 1958 studied the interchange reaction between cystine and N,N'-bis-(2,4-dinitrophenyl)-cystine (bis-DNP cystine) in strong acid solution. Ryle and Sanger found, quite surprisingly, that the interchange was strongly inhibited by the addition of mercaptan in aqueous acid solution. Ryle and Sanger also found that the reaction was strongly acid catalyzed. The product of the reaction, mono-DNP cystine, was formed faster in 10N hydrochloric acid than in 10N sulfuric acid. Benesch and Benesch (5) obtained results which were very similar to those of Ryle and Sanger (70). They found that the reaction was catalyzed by addition of either aliphatic or aromatic sulfonyl chlorides. As the latter have been shown (42) to form sulfenium ions in strongly acidic solution, it

was felt that the reaction must go through a sulfenium ion and the following mechanism was proposed:



They were also able to show that hydrogen peroxide catalyzes the reaction, probably through some mechanism which produces sulfenium ions.

The cystine-bis-DNP cystine interchange reaction system appears to be the only disulfide interchange system which has been studied mechanistically to date. The disulfide interchange system is an example of an organic reaction which goes to a well-defined equilibrium situation. Because of the ease with which the disproportionation occurs it is not easy to analyze for one of the disulfides in the presence of the others without disturbing the equilibrium. Ryle and Sanger (70) found that the unsymmetrical mono-DNP cystine was not extracted from aqueous solution but the bis-DNP cystine was quantitatively removed by two ether extractions. Thus, after quenching with water and extracting twice with ether, they were able to determine concentrations of mono-DNP cystine in the solution by spectrophotometric methods. No rate constants were determined, nor was the order of the reaction with respect to the various reactants

determined. Both groups showed their data in terms of non-linear plots of concentration or percent unsymmetrical disulfide produced versus time.

The equilibrium constant for the disulfide equilibration has been determined for a number of pairs of symmetrical disulfides (6,18,36). Equilibrium data has been determined by use of gas chromatography techniques which are not applicable to determination of rate data. Haraldson, et al, (36) have studied three systems in cyclohexane at two temperatures to determine temperature dependence and effect of steric requirements on the position of equilibrium. Their results are shown in Table X. These authors felt that there was little

Table X

Equilibrium Constants for Some Dialkyl
Disulfide Interchange Reactions *

Reactants	T°C	K_e	
Et ₂ S ₂ +Me ₂ S ₂	25	5.6 ± 0.1	$K_e = \frac{[R-S-S-R]^2}{[RSSR][R'SSR]}$
	60	5.6 ± 0.1	
EtSSMe	25	5.7 ± 0.1	
	60	5.5 ± 0.1	
Et ₂ S ₂ +i-Pr ₂ S ₂	25	4.15 ± 0.15	
Et-SS-i-Pr	25	4.05 ± 0.15	
Et ₂ S ₂ +t-Bu ₂ S ₂	25	24.65 ± 0.1	
	60	24.60 ± 0.1	
Et-S-S-t-Bu	25	24.4 ± 0.1	
	60	24.5 ± 0.1	

* Data from Reference 36.

difference in ΔH° for any of these reactions. The

di-t-butyl disulfide has steric restrictions against rotation about the C-S bonds which are not present in ethyl t-butyl disulfide. Haraldson, et al, state that the difference in equilibrium position is due to a relatively large ΔS° for the ethyl-t-butyl system. Gorin and co-workers (18) have determined the equilibrium constants for a number of disulfide interchanges using both direct determination and an indirect method involving the mercaptan-disulfide interchange reaction. Their results, determined at 25°C in alkaline alcoholic solution, are given in Table XI. These authors felt that in

Table XI

Equilibrium Constants for Some Dialkyl
Disulfide Interchange Reactions *

Reactants	** K _e (direct)	*** K _e (indirect)
n-hexyl ₂ S ₂ +Ø ₂ S ₂	8.1	7.3
n-hexyl ₂ S ₂ +(HOCH ₂ CH ₂) ₂ S ₂	5.0	4.4
n-hexyl ₂ S ₂ +(Et ₂ NCH ₂ CH ₂) ₂ S ₂	5.05	5.4
n-propyl ₂ S ₂ +sec-butyl ₂ S ₂	4.43	4.1
n-butyl ₂ S ₂ +t-butyl ₂ S ₂	-	21

$$K_2 = \frac{[\text{RSSR}]^2}{[\text{RSSR}] [\text{R'SSR}]}$$

* Data from Reference 18.

** Determined from direct mixing of the symmetrical compounds.

*** Indirect-calculated from equilibration of a disulfide with a mercaptan.

the absence of steric effects the position of equilibrium for the disulfide interchange reaction is very close to that which would be predicted statistically, namely, $K=4$. This is in agreement with Birch (6) who concludes that in the absence of steric effects the free energy of the products will be approximately equal to the mean of the free energies of the reactants in the disulfide interchange reaction.

Kice and Morkved (48) recently reported the synthesis of benzyl p-tolyl disulfide. They found that (59) the benzyl methylene protons of this compound showed their nuclear magnetic resonance at somewhat lower field than the methylene protons in benzyl disulfide. Thus the unsymmetrical compound has a singlet at 6.17τ whereas the symmetrical compound absorbs at 6.52τ . It seemed that this difference could be utilized as a method for following the establishment of equilibrium between benzyl disulfide and p-tolyl disulfide in solution. It was felt that this new method could allow one to determine both the equilibrium constant and the rate of reaction for the disulfide interchange reaction without disturbing the equilibrium situation in any manner. This would have important advantages and might permit a more searching study of the mechanism of the disulfide interchange reaction in acid media than had previously been possible.

It was for these reasons that the following work was undertaken.

RESULTS

Equilibrium Studies on the Benzyl Disulfide - p-Tolyl Disulfide Interchange System

An accurate value of the equilibrium constant for the benzyl disulfide-p-tolyl disulfide equilibrium under a variety of conditions is of value for several reasons. First, one could compare the data obtained in acetic acid-sulfuric acid-water solution to the data previously reported (18,36) for other sets of disulfides. Second, knowledge of an accurate value for the equilibrium constant is essential for any subsequent kinetic study of the equilibration rate.

The equilibrium constant was determined at both 0.20 M and 0.40 M sulfuric acid concentration in acetic acid-0.56 M water. The initial concentrations of the two symmetrical disulfides were varied relative to one another, as well as varying the acidity of the solution.

A solution of the two disulfides in acetic acid-sulfuric acid-water was made up in a volumetric flask at room temperature. Aliquots of this solution were then placed in NMR tubes. The NMR tubes were suspended in a constant temperature bath at 70°C and measurements were made after 24 hours and again after at least 48 hours.

The relative concentrations of benzyl p-tolyl disulfide and benzyl disulfide at equilibrium were determined

from the relative areas of the NMR signals due to the methylene protons in the two disulfides.

The results of this study on the equilibrium constant for the benzyl disulfide-p-tolyl disulfide interchange are shown in Table XI. The error in determining the ratio of benzyl p-tolyl disulfide to benzyl disulfide is about 10 to 15% at the maximum. However, since this ratio is squared in the calculation of the equilibrium constant, small errors in $(ArSSR/RSSR)$ cause considerably larger errors in K_e . This presumably is the cause of the considerable scatter in K_e values observed in Table XI. However, the scatter in values appears to be random and there does not seem to be any systematic variation in K_e with time, provided sufficient time is allowed for equilibration to be complete. From the data in Table XI $K_e = 3.2 \pm 0.7$ at 0.40 M sulfuric acid and $K_e = 2.9 \pm 0.7$ at 0.20 M sulfuric acid. There thus appears to be a slight variation in the equilibrium constant with a change in acidity. However, the magnitude of the change is well within the limits of the experimental error. More significant is the fact that these values for K_e are lower than those previously reported (18,36). In fact, this is the first time that the equilibrium constant for a disulfide interchange reaction has been found to be less than the value of 4 predicted on a purely statistical basis.

Table XII
Equilibrium Data for the Benzyl
Disulfide-p-Tolyl Disulfide Interchange

$[RSSR]_{\circ}, M^*$	$[ArSSAr]_{\circ}, M^*$	$[H_2SO_4], M$	$\left(\frac{ArSSR}{RSSR}\right)$ (24 hrs.)	K_e^{**} (24 hrs.)	$\left(\frac{ArSSR}{RSSR}\right)$ (48 hrs.)	K_e^{**} (48 hrs.)
0.100	0.105	0.40	2.00±0.20	3.8±0.7	1.84±0.20	3.2±0.7
0.103	0.100		1.76±0.20	3.2±0.7	1.74±0.20	3.2±0.7
0.15	0.15		1.76±0.20	3.2±0.7	1.70±0.20	3.0±0.7
0.15	0.10		1.27±0.20	3.4±0.7	1.24±0.20	3.2±0.7
0.10	0.15		-	-	2.48±0.20	2.9±0.7
0.10	0.20		-	-	3.33±0.20	3.0±0.7
0.15	0.15	0.20	-	-	1.60±0.20	2.7±0.7
0.15	0.10		2.95±0.20	2.5±0.7	3.24±0.20	3.0±0.7
0.10	0.15		2.37±0.20	2.7±0.7	2.42±0.20	2.8±0.7
0.10	0.20		3.54±0.20	3.1±0.7	3.67±0.20	3.5±0.7
0.10	0.10		1.62±0.20	3.0±0.7	1.66±0.20	3.2±0.7

* R = benzyl, Ar = p-tolyl

**
$$K_e = \frac{[ArSSR]^2}{[RSSR][ArSSAr]}$$

Rate of Equilibration Measurements

The rate of the disulfide disproportionation reaction was followed by the use of the same NMR methods previously mentioned. The solutions were made up with equal amounts of the two symmetrical disulfides so that the simplest possible rate expression would be obtained. Aliquots were then placed in NMR tubes and the NMR tubes were suspended in a rack in the constant temperature bath at 70°C. After allowing the tubes of solution one minute to come to temperature, the first tube was removed (time = t_0) and cooled with ice water to prevent further equilibration. The NMR of the solution was then taken and the ratio of benzyl p-tolyl disulfide to benzyl disulfide was determined in the same manner as was used for determination of the equilibrium constant.

The stoichiometry of the reaction is such that one might expect a first-order dependence of the rate on the concentration of each of the reactants. Frost and Pearson (30, p. 187-188) give the following expression governing the rate of approach to equilibrium of a reversible second-order reaction where the initial concentrations of the two reactants are equal (Equation 20).

$$\ln Q = \ln \left[\frac{x(a-2x_e) + ax_e}{ax_e - ax} \right] = \frac{2a(a-x_e)}{x_e} kt \quad (\text{Eq. 20})$$

$$a = [\text{RSSR}]_0 = [\text{ArSSAr}]_0; \quad x = [\text{RSSR}]_0 - [\text{RSSR}]_t;$$

$$x_e = [\text{RSSR}]_0 - [\text{RSSR}]_e$$

When the data for any given run in the present studies were plotted according to Equation 20, curved, rather than straight, lines were obtained (see Figure 6). Possible causes of this curvature and its significance will be dealt with in the Discussion.

Dependence of Rate on Acidity

Despite the curvature of the data when plotted according to Equation 20 the results of the various runs clearly indicate that the rate of equilibration is strongly dependent on the acidity of the solution. Although no rate constant can be determined from curved plots such as in Figure 6 a qualitative measure of the effect of acid on the rate can hopefully be obtained from a comparison of the acidity to the time, t , required for the reaction to reach a specified value of Q . In the Hammett treatment of acid catalysis (38, p. 121) the data were considered meaningful if $H_0 + \log K \cong \text{Constant}$, indicating that a plot of $\log k$ versus H_0 would have a slope of 1. Since $k \sim \frac{1}{t}$, we are interested in the function $H_0 - \log t_Q$. This function is shown for two values of Q in Table XII.

Table XIII

Relationship of H_o to $\text{Log } t_Q$ for the
Benzyl Disulfide-p-Tolyl Disulfide Interchange

$Q = 2$

t (min.)	H_o	log t	$H_o - \log t$
102	-1.13	2.01	-3.14
37	-1.56	1.57	-3.13
11	-2.08	1.04	-3.12

$Q = 3$

240	-1.13	2.38	-3.51
88	-1.56	1.95	-3.51
26	-2.08	1.42	-3.50

These results indicate that the rate of reaction is linearly related to the acidity of the solution.

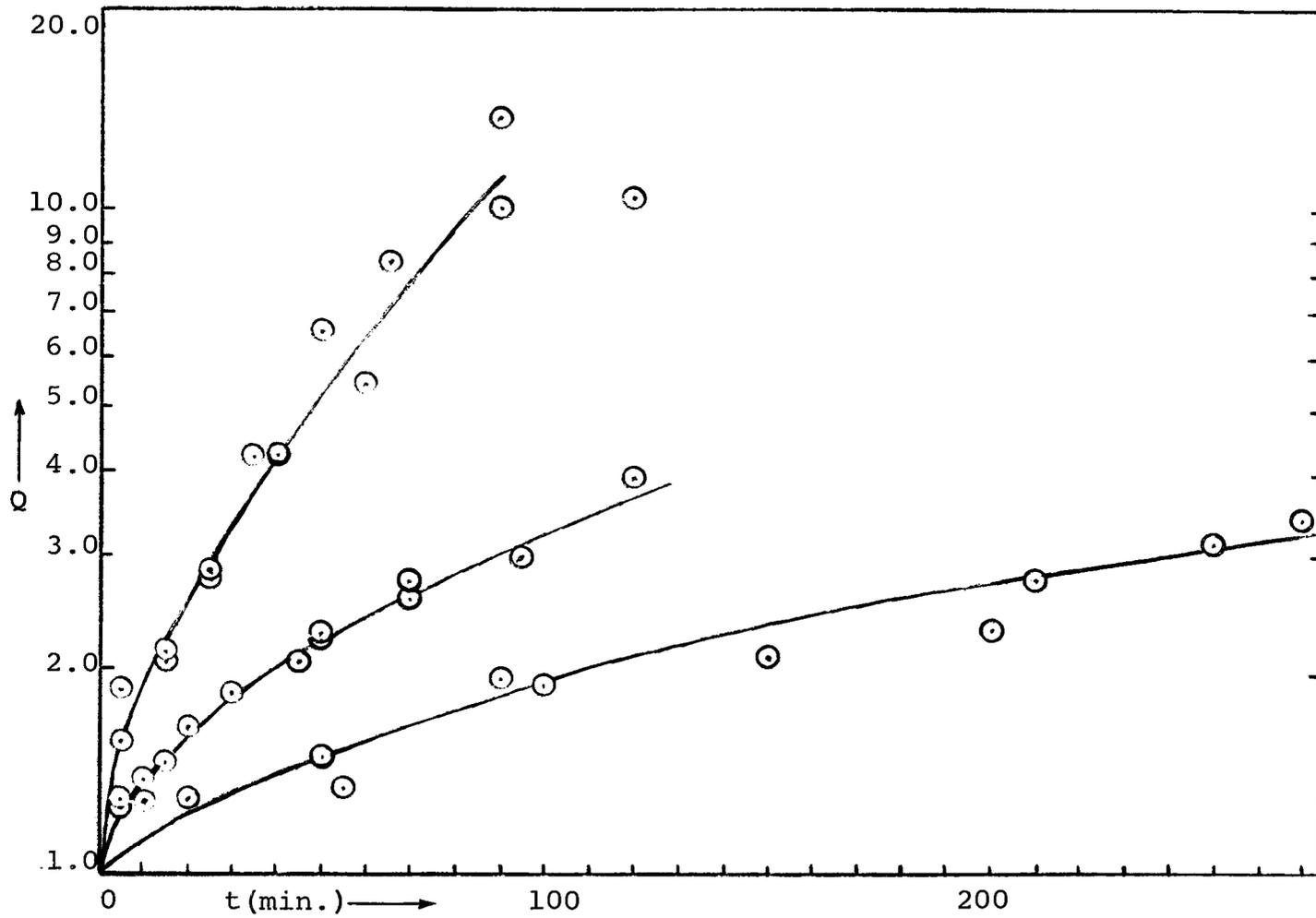


Figure 6. Plot of $\ln Q$ versus t (min.)

DISCUSSION

Equilibrium Constants for the Benzyl Disulfide-
p-Tolyl Disulfide Interchange

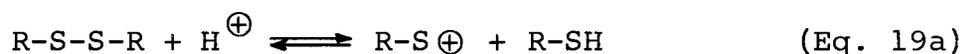
Previous work on other systems (18,36) has shown that most disulfide interchange systems have equilibrium constants of around 4 or larger. In cases where the value of the constant is significantly larger than 4 (the statistically predicted value), the result is quite logically explained on the grounds of steric interactions in one of the symmetrical disulfides (36). Haraldson and co-workers (36) have shown that there is little if any effect of temperature on the position of equilibrium. This indicates that ΔH° for the reaction is very close to zero for all disulfide interchange systems. Any differences in relative equilibrium constants for two disulfide interchange systems must be almost totally due to differences in ΔS° for the reactions. This agrees quite well with what has been observed for systems where steric hindrance to rotation is expected. In a system where the statistically expected equilibrium constant is observed, the free energy of the unsymmetrical compound must very nearly equal the mean of the free energies of the reactants (6).

For the benzyl disulfide-p-tolyl disulfide system in acetic acid-sulfuric acid-water solution the

equilibrium constant is less than what would be predicted by statistical considerations and seems to be slightly dependent on the acidity of the solution. Evidently the mean free energy of the reactants must be lower than the free energy of the unsymmetrical disulfide. This is the first case where the equilibrium constant has been found to be less than 4. One can state that benzyl p-tolyl disulfide is slightly less stable with respect to disproportionation to the symmetrical disulfides than are other unsymmetrical disulfides which have been studied.

Acidity Dependence of the Disulfide Disproportionation

These results for the benzyl p-tolyl disulfide system agree with those obtained by Ryle and Sanger (70) and Benesch and Benesch (5) in that the disulfide interchange reaction is appreciably catalyzed by strong acid. The mechanism proposed by Benesch and Benesch (5) includes a reversible protonation-dissociation (Equation 19a) of one of the disulfides as the first



step in the reaction. The acidity dependence that we have observed is reasonable for this type of a mechanism. However, one should realize that there are many other possible mechanisms for disulfide interchange that could also lead to such a dependence on acidity.

Curvature of Kinetic Plots

There are several reasons why plots of the kinetic data as a reversible second-order reaction might exhibit the downward curvature observed in Figure 6.

First, if there were an impurity present in one of the disulfides which could act as a catalyst for the disulfide interchange but which was somehow slowly destroyed by the acidic reaction solution, the apparent rate of the second-order equilibration would decrease with time.

A second possibility is that during the equilibration small amounts of other sulfur compounds, such as thiols, are produced and that the concentration of these substances gradually rises during the early stages of the reaction until their equilibrium concentration is reached. Should the rate of disulfide equilibration be inversely proportional to the concentration of one of these substances, as has, for example, been claimed to be the case with thiols (5), the second-order rate constant will decrease during the initial stages of the equilibration and will finally level off once the equilibrium concentration of the "retarder" is reached. Since there is some indication that the plots in Figure 6 do become essentially linear in the later stages of the reaction, such an explanation deserves serious consideration.

Finally, a third alternative is that equilibration occurs by both second- and third-order processes, and that the rates of these are comparable under the present conditions. Unfortunately we have not yet been able to derive a satisfactory expression for a second- plus third-order reversible reaction (akin to Equation 20) to test this possibility.

During the present work there was not time to conduct any of the experiments necessary to determine which of these possibilities is the correct one. Further study of the system in the future is therefore indicated.

EXPERIMENTAL

1.) Purification of Symmetrical Disulfides

Benzyl Disulfide (Aldrich Company) was recrystallized from ethyl alcohol until a melting point of 71-72°C was obtained. Literature melting point (37, p. 670), 71-72°C.

p-Tolyl Disulfide (Aldrich Company) was recrystallized from ethyl alcohol, melting point, 48-49°C. Literature melting point (37, p. 1064), 48°C.

2.) Preparation of Stock Solutions

Glacial acetic acid (Baker and Adamson, reagent grade) was refluxed with 10% of its volume of acetic anhydride for 24 hours to remove all traces of water. The acid was then distilled through an Oldershaw column (20 theoretical plates) with only the large center fraction being used for kinetic experiments.

A solution of 0.56 M water in acetic acid was prepared by weighing the necessary amount (10.091 grams) of water into a one-liter volumetric flask and adding purified acetic acid with shaking. The solution was not made up to volume with the purified acetic acid until the water-acetic acid solution had been allowed to equilibrate in a water bath at 20°C for thirty minutes.

Reagent grade sulfuric acid (Baker and Adamson) was

used to prepare 25 ml. of 2 M sulfuric acid-0.56 M water in acetic acid solution. The sulfuric acid was calculated to contain 3.75% water, therefore 5.1016 grams of sulfuric acid and 0.0612 grams of water were weighed into the 25 ml. volumetric flask. All solutions were equilibrated at 20°C for 30 minutes before filling to volume. In this manner the stock solution of sulfuric acid-0.56 M water in acetic acid was made up with five to twenty times the sulfuric acid concentration desired for the various kinetic runs.

3.) Procedure for Determining Equilibrium Constants

The desired amounts of the two symmetrical disulfides were weighed into a 5 ml. volumetric flask and the necessary amount of 3 M sulfuric acid-0.56 M water in acetic acid solution was added with a pipette. The solution was shaken until all of the solid was dissolved, then filled to the mark with acetic acid-0.56 M water. All volumes were made up at room temperature. The solution was then pipetted into NMR tubes which were suspended in a constant temperature bath at 70°C for a period of 24 hours. At this time half of the NMR tubes were removed from the bath, rinsed with acetone and dried with a clean towel. The NMR spectrum of the region between 5.9 and 6.7 τ was taken and several integrals were determined for the areas beneath the two peaks at

6.17 τ and 6.52 τ . From these integrals the average value for the area under each peak was calculated. Using these areas, the ratio of the concentration of benzyl p-tolyl disulfide to the concentration of benzyl disulfide was easily calculated. This ratio was then used to calculate the equilibrium constant as shown in Table XI.

4.) Procedure for Following the Rate of the Benzyl Disulfide p-Tolyl Disulfide Interchange

Equal amounts of each disulfide were used in making up the solutions so that the resulting kinetic data would be in the simplest possible form. One of the symmetrical disulfides was weighed into a 10 ml. volumetric flask and the other symmetrical disulfide was weighed into a small beaker. The desired amount of the 3 M sulfuric acid-0.56 M water in acetic acid solution was pipetted into the volumetric flask and a few ml. of 0.56 M water in acetic acid was added to the beaker. Both solutions were warmed gently on a hot plate until the disulfides went into solution, then both were allowed to cool to room temperature in a desiccator which contained a beaker of 0.56 M water-acetic acid solution. The two solutions were mixed together in the 10 ml. volumetric flask and then filled to the mark with 0.56 M water in acetic acid. Seven or eight NMR tubes were rapidly filled with the

solution and placed in the constant temperature bath at 70°C. The first tube was removed ($t=0$) after the tubes had been in the bath for one minute. No appreciable amount of reaction was found to take place before t_0 when this method for preparing the solutions was used. The ratio of the concentration of benzyl p-tolyl disulfide to the concentration of benzyl disulfide was determined by use of the NMR spectrum in the same manner as was described in the section on equilibrium constants. From this data the concentrations of benzyl disulfide at the various times of tube removal were calculated. These concentrations were then used to calculate the quantity, Q , which is plotted versus t in Figure 6.

BIBLIOGRAPHY

1. Armstrong, Robert T., Julian R. Little and Kenneth W. Doak. Chemistry of sulfur-olefin reactions. Applications to vulcanization. Industrial and Engineering Chemistry 36:628-633. 1944.
2. Arnold, Robert C., Arthur P. Lien and Robert M. Alm. The action of lithium aluminum hydride on organic disulfides. Journal of the American Chemical Society 72:731-733. 1950.
3. Bäder, Erich and Hans Dieter Hermann. Neue Methoden zur Darstellung von Sulfonaminen (IV. Mitteilung: Polymerisationen und Polymerisationskatalystoren). Chemische Berichte 88:41-49. 1955.
4. Bellamy, L. J. The infra-red spectra of complex molecules. 2d ed. London, Methuen, 1962. 425 p.
5. Benesch, Ruth E. and Reinhold Benesch. The mechanism of disulfide exchange in acid solution; role of sulfenium ions. Journal of the American Chemical Society 80:1666-1669. 1958.
6. Birch, S. F., T. V. Cullum and R. A. Dean. The preparation and properties of di-alkyl di- and poly-sulfides. Some disproportionation reactions. Journal of the Institute of Petroleum 39:206-219. 1953.
7. Brand, K. Über o-Azothioanisol und o-Thiodianisidin. (I. Mitteilung über Thiophenole). Berichte der Deutschen Chemischen Gesellschaft 42:3463-3468. 1909.
8. Brederick, Hellmut, Gerhard Brod and Günther Höschele. Die Struktur der Sulfinssäuren. V. Mitteilung über Polymerisationen und Polymerisationskatalysatoren. Chemische Berichte 88:438-444. 1955.
9. Brederick, Hellmut et al. Disproportionierung der Sulfinssäuren und die Struktur der Sulfinssäureanhydride. Angewandte Chemie 70:268-269. 1958.
10. Brintzinger, Herbert and Hans Ellwanger. Synthesen mit α -Chloräthylschwefelchlorid (VII). Mitteilung über organische Schwefelchloride. Chemische Berichte 87:300-314. 1954.

11. Brown, Herbert C. and Y. Okamoto. Electrophilic substituent constants. *Journal of the American Chemical Society* 80:4979-4987. 1958.
12. Bunnett, Joseph F. Kinetics of reactions in moderately concentrated aqueous acids. I. Classification of reactions. *Journal of the American Chemical Society* 83:4956-4967. 1961.
13. Burkhard, R. K. et al. The pK_a 's of aromatic sulfinic acids. *Journal of Organic Chemistry* 24:767-769. 1959.
14. Cahours, A. Ueber einige neue, schwefehaltige Verbindungen des Aethyls und Methyls. *Annalen der Chemie und Pharmacie* 61:91-101. 1847.
15. Cairns, T. L. et al. Gem-dithiols. *Journal of the American Chemical Society* 74:3982-3989. 1952.
16. Calvin, Melvin. Mercaptans and disulfides. Oak Ridge, Tenn., 1954. 39 p. (U.S. Atomic Energy Commission report UCRL-2438).
17. Collin, G. et al. The preparation of aliphatic mercaptans and sulphonic acids of high molecular weight. *Journal of the Society of Chemical Industry* 52:272-275. 1933.
18. Dalman, Gary, John McDermed and George Gorin. Mercaptan-disulfide interchange reactions. II. Some equilibrium constants. *Journal of Organic Chemistry* 29:1480-1484. 1964.
19. Dann, Arthur Thurlby and William Davies. The reactions of nitrosulphonyl chlorides. Part I. The reaction of hydrazine hydrate with o-nitrosulphonyl chlorides. *Journal of the Chemical Society*, 1929, p. 1050-1055.
20. Danson, Joseph. On the identity of bisulphamylic and hyposulphamylic acids. *Journal of the Chemical Society* 3:158-161. 1851.
21. Detoni, S. and D. Hadži. Absorption spectra and structure of sulfinic acids. *Journal of the Chemical Society*, 1955, 3163-3169.

22. Dornow, Alfred. Über die Einwirkung von Alkali auf Arylthiosulfate. Berichte der Deutschen Chemischen Gesellschaft 72:568-570. 1939.
23. Douglass, Irwin B., Frederic T. Martin and Roger Addor. Sulfenyl chloride studies. II. Mono-, di-, and tri-chloromethanesulfenyl chlorides and certain of their derivatives. Journal of Organic Chemistry 16:1297-1302. 1951.
24. Dutcher, James D., John R. Johnson and William F. Bruce. Gliotoxin. VI. The nature of sulfur linkages. Conversion to desthiogliotoxin. Journal of the American Chemical Society 67:1736-1745. 1945.
25. Ferington, T. E. and A. V. Tobolsky. Organic disulfides as initiators of polymerization: tetramethylthiuram disulfide. Journal of the American Chemical Society 77:4510-4512. 1955.
26. Friedmann, E., D. H. Marrian and (Mrs.) I. Simon-Reuss. Antimitotic action of maleimide and related substances. British Journal of Pharmacology and Chemotherapy 4:105-108. 1949.
27. Fries, K., H. Koch and H. Stukenbrock. Zur Kenntnis des Thianthrenes. Justus Liebig's Annalen der Chemie 468:162-201. 1929.
28. Fries, K. and G. Schurmann. Reduktion der sulfinsäuren mittels bromwasserstoffs. Berichte der Deutschen Chemischen Gesellschaft 47:1195-1203. 1914.
29. Fromm, Emil and José de Seixas Palma. Ueber Benzylsulfinsäure und Benzyl-disulfoxyd. Berichte der Deutschen Chemischen Gesellschaft 39:3308-3317. 1906.
30. Frost, Arthur A. and Ralph G. Pearson. Kinetics and mechanism. 2d ed. New York, Wiley, 1961. 405 p.
31. Gilman, Henry (ed.). Organic Synthesis. Coll. Vol. 1. New York, Wiley, 1941. 580 p.
32. Glass, H. B. and E. Emmet Reid. The direct introduction of sulfur into aromatic hydrocarbons. Journal of the American Chemical Society 51:3428-3430. 1929.

33. Gringras, L. and G. Sjöstedt. Potentiometric determination of arylsulfonic acids. *Acta Chemica Scandinavica* 15:433-434. 1961.
34. Gur'yanova, E. N. and Ya. K. Surkin. Dipole moments and vibrational spectra of sulfonic acids. *Zhurnal Fizicheskoi Khimi* 23:105-114. 1949. (Abstracted in *Chemical Abstracts* 43:5245f. 1949.)
35. Hall, Maynard E. Polarographic studies of sulfur compounds in petroleum fractions. *Analytical Chemistry* 25:556-561. 1953.
36. Haraldson, L. et al. Equilibrium studies on the disproportionation reaction between some dialkyl disulfides. *Acta Chemica Scandinavica* 14:1509-1514. 1960.
37. Heilbron, I. M. and H. M. Bunberry. *Dictionary of organic compounds*. Rev. ed. Vol. 1 London, Eyre and Spottiswoode, 1943. 1072 p.
38. Hine, Jack. *Physical organic chemistry*. 2d ed. New York, McGraw-Hill, 1962. 552 p.
39. Horner, Leopold and Ottokar H. Basedow. Zum Mechanismus der Autoxydation der Benzolsulfonsäure. *Justus Liebig's Annalen der Chemie* 612:108-131. 1958.
40. Houben, J. and Hans Doescher. Ueber Hydropinensulfonsäure, Hydropinencarbithiosäure, Thioborneol und Thiocampher. *Berichte der Deutschen Chemischen Gesellschaft* 39:3503-3509. 1906.
41. Kalle, Wilhelm. Über Benzylschweflige Säure. *Annalen der Chemie und Pharmacie* 119:153-164. 1861.
42. Kharasch, Norman, Charles M. Buess and William King. Derivatives of sulfenic acids. XIV. 2,4-dinitrobenzenesulfenium ion. *Journal of the American Chemical Society* 75:6035-6038. 1953.
43. Kice, John L. and Kerry W. Bowers. The mechanism of the disproportionation of sulfenic acids. *Journal of the American Chemical Society* 84:605-610. 1962.

44. _____ . Mechanisms of reactions of sulfinic acids. II. The reaction of p-tolyl disulfide with p-toluenesulfinic acid. Journal of the American Chemical Society 84:2384-2389. 1962.
45. _____ . Mechanisms of reactions of sulfinic acids. III. The reaction of some dialkyl sulfides with p-toluenesulfinic acid. Journal of the American Chemical Society 84:2390-2394. 1962.
46. Kice, John L. and Norman E. Pawlowski. The mechanism of the disproportionation of sulfinic acids. The thermal decomposition of p-toluenesulfinyl p-tolyl sulfone and its reaction with p-toluenesulfinic acid. Journal of Organic Chemistry 28:1162-1163. 1963.
47. _____ . The decomposition of aromatic sulfinyl sulfones (sulfinic anhydrides). The facile homolysis of a sulfur-sulfur bond. Journal of the American Chemical Society 86:4898-4904. 1964.
48. Kice, John L. and Eva H. Morkved. Mechanisms of sulfinic acids. VI. The mechanism of the disulfide-sulfinic acid reaction. Journal of the American Chemical Society 86:2270-2278. 1964.
49. Kice, John L., David C. Hampton and Alvin Fitzgerald. The effect of structure on the rate of disproportionation of aromatic sulfinic acids. Journal of Organic Chemistry 30:882-885. 1965.
50. Kolbe. Vorläufige Notiz über einige neue Derivate des Benzylsulfonchlorids. Annalen der Chemie und Pharmacie 115:352-354. 1860.
51. Lefevre, C. and C. Desgrez. Organische Schwefelen Verbindungen. Comptes rendus 198:1432-1434. 1934.
52. Mannich, C. and Ph. Fresenius. Über den Hauptbestandteil des ätherischen Öles der Asa foetida. Archiv der Pharmazie und Berichte der Deutschen Pharmazeutischen Gesellschaft 274:461-472. 1936.
53. Marvel, C. S. and Rayner S. Johnson. 1-Dodecane-sulfinic acid. Journal of Organic Chemistry 13:822-829. 1948.

54. McAllan, D. T. et al. The preparation and properties of sulfur compounds related to petroleum. I. The dialkyl sulfides and disulfides. *Journal of the American Chemical Society* 73:3627-3632. 1951.
55. Meves, S. Wilhelm. Paraffinic substances containing sulfur. German patent 729,772. Nov. 26, 1942. (Abstracted in *Chemical Abstracts* 38:374. 1944.
56. Milligen, Brian and J. M. Swan. Unsymmetrical disulfides from Bunte salts. *Journal of the Chemical Society*, 1963, p. 6008-6012.
57. Morgan, Marcus S. and Leonard H. Cretcher. A kinetic study of alkylation by ethyl arylsulfonates. *Journal of the American Chemical Society* 70:375-378. 1948.
58. Morin, Pogg. Zweifach Schwefeläthyl. *Annalen der Pharmacie* 32:267-270. 1839.
59. Morkved, Eva Henmo. Mechanisms of reactions of sulfinic acids. I. The sulfide-sulfinic acid reaction. II. The disulfide-sulfinic acid reaction. Ph.D. thesis. Corvallis, Oregon State University, 1964. 193 numb. leaves.
60. Moses, Carrie G. and E. Emmet Reid. The action of sodium on alkyl disulfides. *Journal of the American Chemical Society* 48:476-477. 1926.
61. Muspratt, Sheridan. On the identity of bisulphethylic with hyposulphethylic acids, and bisulphimethylic with hyposulphamethylic acids. *Journal of the Chemical Society* 3:18-23. 1851.
62. Onufrowicz, Stanislas. Ueber Sulfide des β -Naphthols. *Berichte der Deutschen Chemischen Gesellschaft* 23:3355-3373. 1890.
63. Otto, Robert. Ueber Toluolschweflige Säure. *Annalen der Chemie und Pharmacie* 145:10-25. 1868.
64. _____ . Ueber einige Derivate des Benzols und Toluols. *Annalen der Chemie und Pharmacie* 145:317-329. 1868.

65. Otto, Robert and Adalbert Rössing. Synthese von Alkyldisulfiden mit verschiedenen Radikalen. Berichte der Deutschen Chemischen Gesellschaft 19:3132-3138. 1886.
66. . Zur Frage der Tautomerie bei den Sulfinssäuren. Berichte der Deutschen Chemischen Gesellschaft 25:230-233. 1892.
67. Overberger, C. G. and J. J. Godfrey. The sulfinic acid initiated polymerization of methyl methacrylate. Journal of Polymer Science 40:179-201. 1959.
68. Pauley, R. and R. Otto. Zur Kenntniss der Bildung und Constitution des Benzoldisulfoxyds und Paratoluoldisulfoxyds. Berichte der Deutschen Chemischen Gesellschaft 10:2181-2185. 1877.
69. Reid, E. Emmet. Organic chemistry of bivalent sulfur. vol. 3. New York, Chemical Publishing, 1960. 486 p.
70. Ryle, A. P. and F. Sanger. Disulfide interchange reactions. Biochemical Journal 60:535-540. 1955.
71. Schneider, Wilhelm and August Bansa. Über Schwefelzucker und ihre Abkömmlinge, XVIII. Mitteilung: β,β -Diglucosylsulfoxy-sulfide. Berichte der Deutschen Chemischen Gesellschaft 66:1973-1975. 1933.
72. Schöberl, Alfons. Neue Forschungsergebniss an Schafwolle. Angewandte Chemie 54:313-320. 1941.
73. Schönberg, Alexander. Über einige neue Reaktionen organischer Disulfide, sowie über die Oxydation organischer Verbindungen bei Gegenwart von Disulfiden. Berichte der Deutschen Chemischen Gesellschaft 68:163-164. 1935.
74. Strecker, W. Einwirkung von Organomagnesiumverbindungen auf Bortirchlorid, Chloroschwefel, sowie auf das Chlorid und die Ester der schwefligen Säure. Berichte der Deutschen Chemischen Gesellschaft 43:1131-1136. 1910.

75. Taft, Robert W., Jr. A precise correlation of nuclear magnetic shielding in m- and p-substituted fluorobenzene by inductive and resonance parameters from reactivity. *Journal of the American Chemical Society* 79:1045-1049. 1957.
76. Truce, William E. and Anna M. Murphy. The preparation of sulfinic acids. *Chemical Reviews* 48:69-124. 1951.
77. Von Braun, Julius and Karl Weissbach. Zur Kenntnis der organischen Sulfon- und Sulfinsäuren. *Berichte der Deutschen Chemischen Gesellschaft* 63:2836-2847. 1930.
78. Wawzonek, Stanley. Organic polarography. *Analytical Chemistry* 21:61-62. 1949.
79. Zeise, W. C. Ueber das Mercaptan. *Annalen der Pharmacie* 11:1-10. 1834.
80. Zincke, Th. Uber Arylschwefelchloride. *Justus Liebig's Annalen der Chemie* 400:1-2. 1913.

APPENDICES

APPENDIX I

Table XIV

Disproportionation of Arylsulfinic Acids

Compound: p-Bromobenzenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 80.3°C

<u>Run 1</u>		
t(hr.)	M	1/M
0	0.0964	10.37
0.333	0.0877	11.41
0.833	0.0755	13.25
1.500	0.0640	15.62
2.330	0.0536	18.69
3.330	0.0460	21.75

$$k_2 = 1.01 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

<u>Run 2</u>		
t(hr.)	M	1/M
0	0.1000	10.00
1.667	0.0613	16.33
3.500	0.0441	22.70
5.500	0.0320	31.25
6.000	0.0307	32.58
6.500	0.0293	34.10
8.000	0.0260	38.50
9.250	0.0247	40.50

$$k_2 = 1.04 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: p-Methoxybenzenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 80.3°C

<u>Run 3</u>		
t(min.)	M	1/M
0	0.0532	19.8
4.00	0.0468	21.4
7.75	0.0428	23.3
12.25	0.0368	27.2
16.75	0.0340	29.4
21.25	0.0304	32.9
26.00	0.0279	35.7
30.50	0.0260	38.5

$$k_2 = 10.78 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

<u>Run 4</u>		
t(min.)	M	1/M
0	0.0614	16.3
6.50	0.0504	19.8
13.25	0.0407	24.6
20.50	0.0348	28.7
26.50	0.0317	31.7
32.25	0.0272	36.8
41.00	0.0240	41.7

$$k_2 = 10.66 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: p-Nitrobenzenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 80.3°C

<u>Run 5</u>		
t (hr.)	M	1/M
0	0.0911	10.98
1.10	0.0740	13.51
2.23	0.0703	14.25
3.47	0.0658	15.2
4.13	0.0623	16.1
5.12	0.0558	17.9
6.92	0.0500	20.0

$$k_2 = 3.5 \times 10^{-4} \text{ M}^{-1} \text{ sec}^{-1}$$

<u>Run 6</u>		
t (hr.)	M	1/M
0	0.0900	11.1
1.166	0.0810	12.4
2.18	0.0738	13.6
3.00	0.0710	14.1
4.00	0.0614	16.3
5.18	0.0567	17.7
6.16	0.0558	17.9
7.76	0.0498	20.1

$$k_2 = 3.49 \times 10^{-4} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: 2-Naphthalenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 80.3°C

<u>Run 7</u>		
t (hr.)	M	1/M
0	0.0768	13.05
0.317	0.0488	20.5
0.550	0.0376	26.6
0.735	0.0317	31.7
0.917	0.0276	36.2
1.150	0.0240	41.7

$$k_2 = 7.0 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

<u>Run 8</u>		
t (hr.)	M	1/M
0	0.0911	10.98
0.267	0.0592	16.9
0.467	0.0440	22.7
0.650	0.0364	27.4
0.885	0.0300	33.3
1.100	0.0260	38.5
1.265	0.0233	42.1

$$k_2 = 6.9 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: 2-Naphthalenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 70.2°C

Run 9		
t (hr.)	M	1/M
0	0.0852	11.72
0.670	0.0558	17.9
1.150	0.0524	19.1
1.500	0.0428	23.3
2.120	0.0372	26.8
2.620	0.0313	31.8
3.130	0.0256	39.1
3.700	0.0243	41.2
3.930	0.0231	43.3
4.250	0.0224	44.7
4.520	0.0217	46.1

$$k_2 = 2.2 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Run 10		
t (hr.)	M	1/M
0	0.0983	10.2
0.317	0.0776	12.9
0.800	0.0617	16.2
1.485	0.0456	21.9
1.885	0.0408	24.5
2.150	0.0364	27.5
2.630	0.0320	31.2
3.200	0.0283	35.2

$$k_2 = 2.2 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: p-Toluenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 59.6°C

Run 11		
t (hr.)	M	1/M
0	0.1058	9.45
1.618	0.0848	11.78
3.640	0.0672	14.85
5.250	0.0576	17.35
7.030	0.0497	20.2
11.650	0.0360	27.8
15.300	0.0297	33.7
19.350	0.0228	43.8

$$k_2 = 0.428 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Run 12		
t (hr.)	M	1/M
0	0.1070	9.35
2.380	0.0788	12.7
4.970	0.0603	16.6
7.490	0.0488	20.5
10.000	0.0408	24.5
12.350	0.0348	28.7
14.900	0.0303	32.9
17.650	0.0263	37.8

$$k_2 = 0.434 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: p-Toluenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 69.7°C

<u>Run 13</u>		
t (hr.)	M	1/M
0	0.0920	10.85
0.450	0.0818	12.22
1.150	0.0640	15.82
1.920	0.0500	20.00
3.220	0.0380	26.2
4.850	0.0284	35.2
6.120	0.0236	42.3

$$k_2 = 1.43 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

<u>Run 14</u>		
t (hr.)	M	1/M
0	0.0994	10.08
1.135	0.0618	16.2
2.130	0.0484	20.7
3.550	0.0353	28.4
4.880	0.0280	35.7
6.680	0.0220	45.5

$$k_2 = 1.44 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: p-Toluenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 80.3°C

<u>Run 15</u>		
t (hr.)	M	1/M
0	0.0818	12.24
0.300	0.0595	16.8
0.583	0.0472	21.2
0.917	0.0377	26.6
1.215	0.0317	32.1
1.550	0.0272	36.7
1.850	0.0237	42.3
2.020	0.0223	44.7

$$k_2 = 4.52 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

<u>Run 16</u>		
t (hr.)	M	1/M
0	0.0872	11.47
0.350	0.0593	16.9
0.617	0.0469	21.3
0.816	0.0408	24.5
1.050	0.0352	28.4
1.400	0.0304	32.9
1.515	0.0272	36.7
1.870	0.0240	41.7

$$k_2 = 4.48 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: p-Toluenesulfinic Acid
 Solvent: Acetic Acid-0.60 M Sulfuric Acid-0.56 M
 Water
 Temperature: 89.4°C

<u>Run 17</u>		
t (min.)	M	1/M
0	0.0768	13.00
5.50	0.0617	16.22
10.50	0.0513	19.50
15.75	0.0424	23.6
20.00	0.0380	26.3
25.50	0.0328	30.5
31.50	0.0313	32.8
36.50	0.0272	36.7

$$k_2 = 11.9 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

<u>Run 18</u>		
t (min.)	M	1/M
0	0.0838	11.92
6.25	0.0636	15.70
12.00	0.0504	19.85
17.50	0.0417	24.0
23.00	0.0360	27.8
28.75	0.0312	32.1
39.75	0.0257	39.1

$$k_2 = 12.0 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

Compound: p-Toluenesulfinic Acid
 Solvent: Acetic Acid-0.30 M Sulfuric Acid-0.56 M
 Water
 Temperature: 69.7°C

Temperature: 80.3°C

<u>Run 19</u>		
t (hr.)	M	1/M
0	0.1220	8.20
0.650	0.0980	10.20
1.715	0.0712	14.05
2.600	0.0578	17.35
4.150	0.0428	23.3
5.720	0.0348	28.8
7.330	0.0283	35.2
8.650	0.0243	41.0

$$k_2 = 1.03 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

<u>Run 20</u>		
t (hr.)	M	1/M
0	0.0808	12.4
0.500	0.0525	19.1
0.868	0.0420	23.8
1.150	0.0372	26.9
1.815	0.0280	35.7
2.050	0.0260	38.4
2.250	0.0243	41.0
2.470	0.0228	43.8

$$k_2 = 3.2 \times 10^{-3} \text{ M}^{-1} \text{ sec}^{-1}$$

APPENDIX II

Table XV

Disproportionation of Symmetrical Disulfides

$[\text{p-tolyl disulfide}]_0 = [\text{benzyl disulfide}]_0 = 0.10 \text{ M}$
 Solvent: Acetic Acid-Sulfuric Acid-0.56 M water.
 Temperature: 69.7°C.

Run 1			
$[\text{H}_2\text{SO}_4] = 0.10 \text{ M}$			
t (min.)	(ArSSR/RSSR) *	[RSSR] **	Q ***
0	-	0.1000	1.000
5	0.1703	0.0922	1.251
20	0.1887	0.0913	1.286
50	0.3022	0.0869	1.491
90	0.5010	0.0800	1.974
150	0.5470	0.0784	2.128
210	0.721	0.0736	2.772
270	0.839	0.0705	3.426
∞ ****	1.528	0.0567	-

Run 2			
$[\text{H}_2\text{SO}_4] = 0.10 \text{ M}$			
t (min.)	(ArSSR/RSSR)	[RSSR]	Q
0	-	0.1000	1.000
10	0.1850	0.0916	1.243
55	0.2441	0.0892	1.337
100	0.5220	0.0793	1.906
150	0.608	0.0767	2.133
200	0.704	0.0750	2.324
250	0.870	0.0697	3.153
300	0.932	0.0682	3.51
360	1.036	0.0659	4.22
∞	1.611	0.0544	-

* Ar- = p-tolyl- , R- = benzyl- .

** $[\text{RSSR}]_i$ calculated from: $[\text{RSSR}]_i = \frac{[\text{RSSR}]_0 + [\text{ArSSAr}]_0}{2 + (\text{ArSSR}/\text{RSSR})_i}$

*** $Q = \frac{x(a-2x_e) + ax_e}{ax_e - ax}$ from $\ln \left[\frac{x(a-2x_e) + ax_e}{ax_e - ax} \right] = \frac{2a(a-x_e)}{x_e} kt$

where $a = [\text{RSSR}]_0 = [\text{ArSSAr}]_0$; $x = [\text{RSSR}]_0 - [\text{RSSR}]_i$;

$x_e = [\text{RSSR}]_0 - [\text{RSSR}]_e$ from Reference 24, p. 188-189.

**** t_∞ taken as 24 hours.

Run 3

$$[\text{H}_2\text{SO}_4] = 0.20 \text{ M}$$

t (min.)	(ArSSR/RSSR)	[RSSR]	Q
0	0	0.1005	1.000
5	0.1982	0.0915	1.254
10	0.2728	0.0885	1.368
20	0.433	0.0827	1.654
35	0.620	0.0768	2.085
50	0.662	0.0755	2.225
70	0.803	0.0713	2.765
∞	1.750	0.0536	-

Run 4

$$[\text{H}_2\text{SO}_4] = 0.20 \text{ M}$$

t (min.)	(ArSSR/RSSR)	[RSSR]	Q
0	-	0.1005	1.000
5	0.2185	0.0907	1.283
15	0.3275	0.0864	1.463
30	0.529	0.0795	1.867
50	0.683	0.0749	2.293
70	0.772	0.0726	2.574
95	0.872	0.0701	2.975
120	1.091	0.0651	3.960
∞	1.700	0.0537	-

Run 5

$$[\text{H}_2\text{SO}_4] = 0.40 \text{ M}$$

t (min.)	(ArSSR/RSSR)	[RSSR]	Q
0	-	0.1000	1.000
5	0.3785	0.0842	1.573
15	0.613	0.0766	2.138
25	0.800	0.0714	2.813
40	1.041	0.0658	4.225
60	1.156	0.0633	5.43
90	1.381	0.0592	10.04
120	1.384	0.0590	10.47
∞	1.682	0.0543	-

Run 6

$$[\text{H}_2\text{SO}_4] = 0.40 \text{ M}$$

t (min.)	(ArSSR/RSSR)	[RSSR]	Q
0	-	0.1000	1.000
5	0.518	0.0794	1.885
15	0.592	0.0772	2.074
25	0.793	0.0717	2.75
35	1.043	0.0657	4.23
50	1.244	0.0617	6.54
65	1.324	0.0602	8.17
80	1.453	0.0578	13.68
100	1.629	0.0552	(49.5)
∞	1.692	0.0542	-