J. Org. Chem. **1981**, *46*, 2691-2694

Synthesis and Rate of Hydrolysis of an Aromatic Sulfinic Anhydride: Benzene-*o*-disulfinic Anhydride

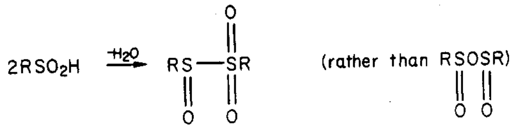
John L. Kice* and Shih-tang Liao

Department of Chemistry, Texas Tech University, Lubbock, Texas 79409

Received January 14, 1981

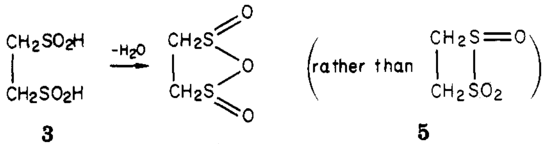
The compound previously (ref 4) thought to be benzene-*o*-disulfinic acid (6) has been shown to be, in fact, the corresponding sulfinic anhydride, 7, the first reported example of an aryl sulfinic anhydride. The kinetics of the hydrolysis of 7 in aqueous dioxane has been determined and compared with those for the hydrolyses of two aromatic sulfinyl sulfones, 8 and 9. Sulfinic anhydride 7 hydrolyzes 4000 times faster than 8 and 200 times faster then 9; ΔH^{*} for its hydrolysis is 6 kcal/mol less than that for 9. The significantly lower ΔH^{*} for the sulfinic anhydride is consistent with other indicators, suggesting that sulfinic anhydrides are normally substantially less stable thermodynamically than the isomeric sulfinyl sulfones.

Dehydration of sulfinic acids (eq 1) normally leads to the formation of a sulfinyl sulfone (1) rather than the isomeric sulfinic anhydride (2).[1] However, among alkanesulfinic acids two cases are known[2] where formation of the sulfinic anhydride is preferred. One of these is ethane-1,2-disulfinic acid (3).[2b] In this instance the strain associated with the four-membered ring in sulfinyl sulfone 5 is apparently sufficient to cause formation of the five-membered sulfinic anhydride 4 to be favored.



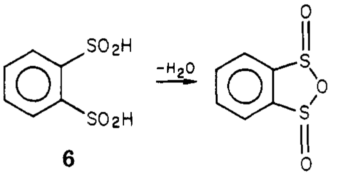
(1)

1 2



3 4 5

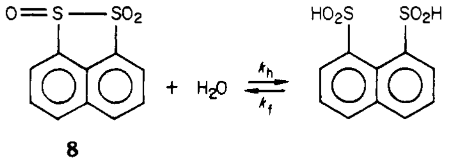
Given the behavior of 3, it seemed likely that loss of water from benzene-*o*-disulfinic acid (6) would lead to benzene-*o*-disulfinic anhydride (7) rather than to the isomeric sulfinyl sulfone, thereby affording the first example of an aromatic sulfinic anhydride.



(2)

6 7

The rate of hydrolysis (k_h, eq 3) of the cyclic, five-membered sulfinyl sulfone 8 is not too different (about 20 times slower) from the rate of hydrolysis of phenyl benzenesulfinyl sulfone, PhS(O)SO$_2$Ph (9),[3a] even though the equilibrium constant ($K_{eq} = k_f/k_h$) for the formation of 8 is many orders of magnitude larger[3b] than the equilibrium constant for the formation of 9 from two molecules of benzenesulfinic acid. This similarity in the rates of hydrolysis of 8 and 9 suggested that, if 7 could be prepared, the kinetic behavior of its hydrolysis might provide a reasonable model for the kinetics of hydrolysis of the acyclic aromatic sulfinic anhydride PhS(O)OS(O)Ph, a compound that itself has not as yet ever been prepared. Comparison of the kinetics of hydrolysis of 7 with those of aromatic sulfinyl sulfones like 8 and 9 under the same conditions should thus be interesting and informative.

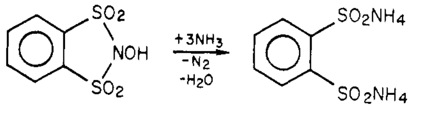


(3)

8

Results

Preparation of Benzene-*o*-disulfinic Anhydride. A compound (mp 113–115 °C) assigned the structure benzene-*o*-disulfinic acid (6) has been reported by Hendrickson et al.[4] We repeated the preparation of this substance, using a slightly modified route. We first synthesized benzene-*o*-disulfonhydroxyimide[4,5] (10) and then used the known[4,5] reaction (eq 4) of 10 with ammonia to prepare a concentrated aqueous solution of diammonium benzene-*o*-disulfinate (11). When this solution of 11 was acidified with 30% sulfuric acid and extracted with methylene chloride, and the extract dried and evaporated, we obtained a compound having a slightly higher melting point (122–123 °C) than that given by Hendrickson and co-workers[4] but whose infrared spectrum agreed with that which they had reported. Because of the correspondence in the infrared spectra we believe that, despite the difference in melting point, our compound and that isolated by Hendrickson and co-workers are, in fact, the same.



(4)

10 11

Further examination indicated, however, that the substance is not the disulfinic acid (6), as previously supposed,[4] but is rather actually benzene-*o*-disulfinic anhydride (7). The experimental evidence leading to this conclusion is as follows. (1) Addition of 10% by volume of 0.10 N aqueous perchloric acid to a solution of the compound (5 × 10^{-4} M) in *anhydrous* dioxane leads to a significant change in the UV spectrum of the solution. This change,

---

[1] Bredereck, H.; Wagner, A.; Beck, H.; Klein, R. J. *Chem. Ber.* **1960**, *93*, 2736.

[2] (a) Kice, J. L.; Ikura, K. *J. Am. Chem. Soc.* **1968**, *90*, 7378. (b) Meuller, W. H.; Dines M. B. *Chem. Commun.* **1969**, 1205.

[3] (a) Chau, M. M.; Kice, J. L. *J. Org. Chem.* **1977**, *42*, 3265. (b) Kice, J. L.; Margolis, H. C. *Ibid.* **1975**, *40*, 3623.

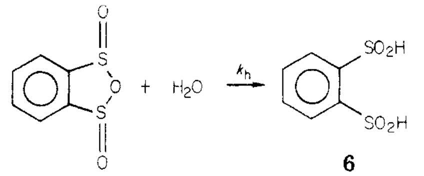
[4] Hendrickson, J. B.; Okano, S.; Bloom, R. K. *J. Org. Chem.* **1969**, *34*, 3434.

[5] Smiles, S.; Hertley, W. R. H. *J. Chem. Soc.* **1926**, 1821.

0022-3263/81/1946-2691$01.25/0

while rapid, is not instantaneous and can, in fact, be followed kinetically. If the compound were **6**, any change in UV spectrum that would occur upon addition of a small amount of dilute aqueous acid should be instantaneous. On the other hand, if the compound is **7** the change in the UV spectrum with time after addition of aqueous acid is easily understood; what is being observed is the hydrolysis of **7** to **6**. (2) The infrared spectrum (KBr) of the substance exhibits a strong band at 1120–1140 cm^{-1}. The absorption for the sulfinyl group in aromatic sulfinic acids is normally found at somewhat lower frequency than this, ~1090 cm^{-1}.[6] On the other hand, in the sulfinic anhydride *t*-BuS(O)OS(O)Bu-*t* the stretching frequency for the sulfinyl groups is 1136 cm^{-1}.[2a] The location of the strong infrared band associated with the sulfinyl groups in the compound is thus in better accord with the sulfinic anhydride structure (**7**) than with the disulfinic acid structure (**6**). (3) The NMR ($CDCl_3$) spectrum for the compound consists of an AA′BB′ pattern located between δ 7.88 and 8.10, assigned to the aromatic protons in the molecule. There is no signal further downfield (δ >9.0) such as would be expected if SO_2H protons were also present. The NMR spectra of two typical aromatic sulfinic acids, benzenesufinic and *p*-toluenesulfinic acids, were measured under the same conditions and in each case the SO_2H protons were plainly visible ($PhSO_2H$, δ 9.9; *p*-$MeC_6H_4SO_2H$, δ 9.2). The fact that no signal that could be ascribed to the SO_2H protons of **6** can be found in the NMR spectrum of the compound thus also suggests that the compound is the sulfinic anhydride **7** rather than the disulfinic acid **6**.

Kinetics of Hydrolysis of Benzene-*o*-disulfinic Anhydride (7). The kinetics of the hydrolysis of **7** to **6** (eq 5) was studied in 90% dioxane–10% H_2O (v/v), using stopped-flow spectrophotometry to follow the reaction. The disappearance of **7** followed good first-order kinetics in all cases. Table I lists the first-order rate constants (k_h) for the hydrolysis of **7** under the various reaction conditions.



(5)

7 **6**

In the region [H$^+$] = 0.01–0.10 M, k_h increases with increasing acid concentration. However, this increase is apparently not due to acid catalysis but rather to a salt effect, since addition of an equivalent amount of lithium perchlorate has the same effect on the rate. Above [H$^+$] = 0.30 M the rate begins to decrase with increasing acid concentration. This is probably due to the fact that at these relatively high acidities the decrease in a_{H_2O} caused by increasing [H$^+$] more than compensates for any increase in rate that would be associated with increasing ionic strength.

Since acid catalysis is not important, the rate constant for the hydrolysis in the presence of 0.01 M added $HClO_4$ can be taken as equal to the rate of spontaneous hydrolysis of **7**. In 95% dioxane–5% water the rate is about 3 times slower than in 90% dioxane–10% water. Runs in 90% dioxane–10% D_2O indicate that the spontaneous hydrolysis of **7** exhibits a sizeable solvent isotopoe effect, k_{H_2O}/k_{D_2O} = 2.3. From the variation of k_h with temperature $\Delta H^\ddagger$ = 4.3 kcal/mol and $\Delta S^\ddagger$ = −49.7 eu.

Table. I. Kinetics of Hydrolysis of Benzene-*o*-disulfinic Anhydride in Aqueous Dioxane

temp, °C	C_{H_2O}, %[a]	C_{LiClO_4}, M	C_{HClO_4}, M	$k_h \times 10^2$, s^{-1} [b]
25.0	5	0.00	0.01	1.9
	10	0.00	0.01	6.1
			0.04	7.4
			0.06	8.0
			0.10	9.7
			0.30	9.4
			0.50	7.5
		0.09	0.01	9.5
	10 (D_2O)	0.00	0.01 ($DClO_4$)	2.7
34.8	10	0.00	0.01	7.5
35.0	10	0.00	0.01	8.1
43.7	10	0.00	0.01	9.8

[a] Concentration of water (v/v). [b] In most cases the listed rate constants represent the average of several runs. Rate constants were reproducible to ±5%. Initial concentration of **7**, 2.5 × 10^{-4} M in all runs.

Table II. Kinetics of Hydrolysis of Phenyl Benzenesulfinyl Sulfone in 90% Dioxane (v/v)

temp, °C	C_{LiClO_4}, M	C_{HClO_4}, M	$k_h' \times 10^2$, s^{-1} [a]
25.0	0.00	0.01	0.028
		0.02	0.030
		0.04	0.032
		0.06	0.033
		0.10	0.038
		0.30	0.040
		0.50	0.032
	0.09	0.01	0.038
	0.00	0.01	0.016 (D_2O)
35.0	0.00	0.01	0.052
45.0	0.00	0.01	0.092

[a] In most cases the listed rate constants represent the average of several runs. Rate constants were reproducible to ±3%. Initial concentration of **9** 1 × 10^{-4} M in all runs.

To provide a comparison of the hydrolysis behavior of **7** with that of aryl sulfinyl sulfone **9** under the same conditions the kinetics of the hydrolysis of **9**, which had previously been investigated only in more aqueous dioxane (20–80% dioxane), was determined in 90% dioxane. The results are summarized in Table II. The rate constants show the same type of variation with perchloric acid and lithium perchlorate concentration as the hydrolysis of **7**. The solvent isotope effect, k_{H_2O}/k_{D_2O} = 1.7, is somewhat smaller than that for **7**. The activation parameters are $\Delta H^\ddagger$ = 10.3 kcal/mol and $\Delta S^\ddagger$ = −40.2 eu.

Discussion

The compound previously[4] assigned the structure of benzene-*o*-disulfinic acid (**6**) is actually the corresponding sulfinic anhydride (**7**) and represents the first known example of an aromatic sulfinic anhydride. That **7** should be isolated rather than disulfinic acid **6** is not surprising given the behavior of several other aromatic disulfinic acids. Thus acidification of solutions of either naphthalene-1,8-disulfinate[3b] or biphenyl-2,2′-disulfinate[3a] leads to the separation not of the corresponding disulfinic acid but rather of the cyclic sulfinyl sulfone, arising from the dehydration of the disulfinic acid (**8** in the case of naphthalene-1,8-disulfinate). In the present case we believe that while acidification of an aqueous solution of **11** undoubtedly forms **6**, and that while this is presumably what is

---

[6] Bellamy, B. J., "Organic Sulfur Compounds"; Kharasch, N., Ed.; Pergammon Press: London, 1961; Vol. 1, p 47.

[7] Kice, J. L.; Guaraldi, G. *J. Am. Chem. Soc.* **1967**, *89*, 4113.

extracted from the acidified solution by methylene chloride, the **6** $\rightleftarrows$ **7** + H_2O equilibrium is sufficiently mobile so that when the methylene chlroide solution is dried **6** undergoes facile conversion to **7**.

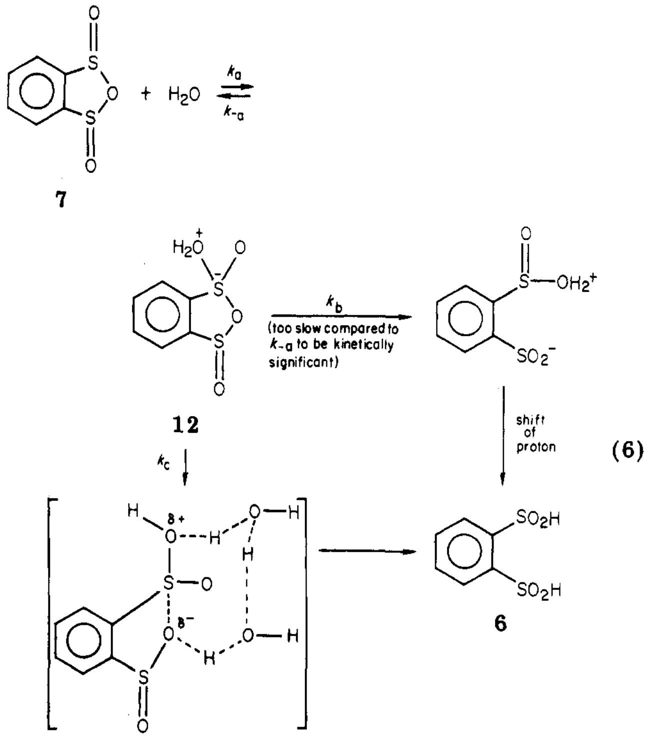
Comparison of the spontaneous hydrolysis of **7** in 90% dioxane–10% H_2O with that of a typical aromatic sulfinyl sulfone (**9**) reveals that sulfinic anhydride **7** hydrolyzes about 200 times faster than **9**. Since the rate of hydrolysis of the cyclic, five-membered aromatic sulfinyl sulfone **8** has been found[3b] to be about 1/20th as fast as that of **9**, it appears that a cyclic five-membered aromatic sulfinic anhydride like **7** hydrolyzes about 4000 times faster than a cyclic five-membered aromatic sulfinyl sulfone (**8**).

The $\Delta H^{\ddagger}$ for the spontaneous hydrolysis of **7** is very small (4.3 kcal/mol) and is 6 kcal/mol less than $\Delta H^{\ddagger}$ for the spontaneous hydrolysis of **9**. It is only because the $\Delta S^{\ddagger}$ for the hydrolysis of **7** is 9 eu more negative than that for **9** that the ratio of the rates of hydrolysis of **7** and **9** is only 200. Were the relative rates of hydrolysis of **7** and **9** determined solely by $\Delta(\Delta H^{\ddagger})$ the sulfinic anhydride would hydrolyze 2.2×10^4 faster than the sulfinyl sulfone.

As outlined in the introduction, there is reason to feel that the kinetic behavior of the hydrolysis of **7** may provide a reasonable model for what might be expected for the hydrolysis behavior of the acylic sulfinic anhydride PhS(O)OS(O)Ph, a compond that has not so far been prepared, presumably because of the significantly greater thermodynamic stability of the isomeric sulfinyl sulfone, PhS(O)SO_2Ph (**9**). The fact the $\Delta H^{\ddagger}$ for the hydrolysis of **7** is 6 kcal/mol less than that for **9** therefore suggests that $\Delta H^{\ddagger}$ for the hydrolysis of PhS(O)OS(O)Ph would be considerably smaller than that for **9**. Since hydrolysis of PhS(O)OS(O)Ph and **9** leads to identical products (2PhSO_2H), such a difference in $\Delta H^{\ddagger}$ is certainly consistent with the view that the sulfinic anhydride PhS(O)OS(O)Ph should be significantly less stable thermodynamically than the isomeric sulfinyl sulfone, **9**.

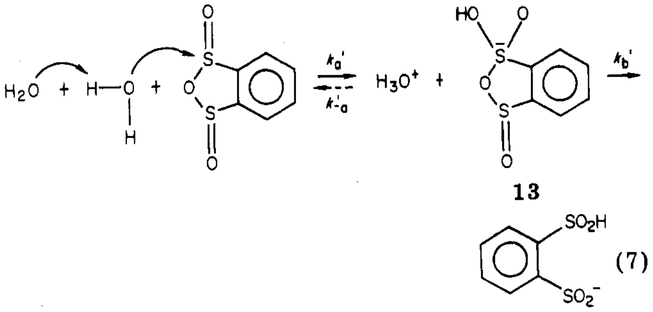
In 90% dioxane containing 0.01 M $HClO_4$ the hydrolysis of **7** to **6** appears to go to completion; i.e., in acidic aqueous dioxane at equilibrium [**7**]/[**6**] is very small. In contrast, the hydrolysis of cyclic sulfinyl sulfone **8** (eq 3) does not go to completion in 60% dioxane containing 0.01 M $HClO_4$,[3b] the ratio [**8**]/[disulfinic acid] being ~3 under these conditions. The fact that the equilibrium constants for eq 3 and 5 are so different is reasonable, however, if a sulfinic anhydride structure is inherently less stable thermodynamically by ~6 kcal/mol than an analogous sulfinyl sulfone.

The size of the solvent isotope effect ($k_{H_2O}/k_{D_2O} = 2.3$) indicates that a proton transfer is part of the rate-determining step of the spontaneous hydrolysis of **7**.[8] The manner by which this proton transfer assists the hydrolysis of **7** is probably analogous to what has been suggested[10] in the mechanism of spontaneous hydrolysis of sulfinyl sulfones. Thus the suggestion is that in intermediate **12** (which would result from attack of water on one of the sulfinyl groups in **7**) H_2O is a so much better leaving group than SO_2^- that step k_b is too slow relative to k_{-a} for the path from **12** via k_b to be of adequate kinetic importance. Either the departure of SO_2^- must be assisted by the transfer of a proton to it coincident with the cleavage of the O–S bond (step k_c, eq 6) or, if SO_2^- is to depart as the anion, one must



(6)

bypass the formation of **12** and follow an alternate mechanism (eq 7) where one water molecule acts as a base to

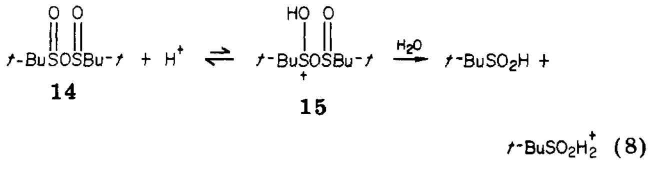


(7)

remove a proton from another water synchronous with its attack on the sulfinyl group, and intermediate **13** is formed; in **13** departure of SO_2^- (step k_b') is much faster than the H_3O^+-catalyzed reversion of **13** to **7** (step k_{-a}'). Which of these two—eq 6, step k_c, or eq 7, step k_a' is the actual correct representation for the rate-determining step of the spontaneous hydrolysis of **7** cannot yet be established.

In 90% dioxane acid catalysis of the hydrolysis of **7** or **9** is not observed, the increase in rate with increasing acid concentration in the range [H^+] = 0.01–0.10 M being due to a salt effect. While the hydrolysis of alkyl sulfinyl sulfones in 60% dioxane also does not exhibit acid catalysis, that of an alkanesulfinic anhydride, *t*-BuS(O)OS(O)Bu-*t* (**14**), does, the rate constant for its hydrolysis (k_h) being given by $k_h = k_0 + k_H[H^+]$.[2a] One therefore needs to explain why an acid-catalyzed term is detected in the hydrolysis of alkanesulfinic anhydride **14** but not in the hydrolysis of arenesulfinic anhydride **7**.

The acid-catalyzed term in the hydrolysis of **14**, which at [H^+] = 0.10 M is responsible for only 25% of the total rate of hydrolysis, was shown[2a] to result from attack of water on the protonated sulfinic anhydride (**15**), eq 8. A



$$t\text{-BuS(O)OS(O)Bu-}t + H^+ \rightleftarrows t\text{-BuS(OH)}\overset{+}{\text{O}}\text{S(O)Bu-}t \xrightarrow{H_2O} t\text{-BuSO}_2\text{H} + t\text{-BuSO}_2\text{H}_2^+ \quad (8)$$

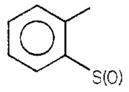
---

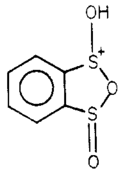
[8] In the spontaneous hydrolyses of carboxylic acid derivatives a large k_{H_2O}/k_{D_2O} and negative $\Delta S^{\ddagger}$ are considered[9] as indicating a mechanism where a proton transfer is part of the rate-determining step and where there is substantial reorganization of water molecules about the transition state.

[9] Johnson, S. L. *Adv. Phys. Org. Chem.* **1967**, *5*, 67.

[10] Kice, J. L. *Adv. Phys. Org. Chem.* **1980**, *17*, 65.

sulfinyl group in 7 should be much less basic than one in 14 since represents a strongly electron-withdrawing group as compared to *t*-Bu.[11] Therefore, at a given acidity the fraction of 7 present in the protonated form (16) should be much smaller than the fraction of 14 present as 15. Apparently the effect is large enough so that an acid-catalyzed pathway involving reaction of water with 16 is not competitive kinetically with the spontaneous hydrolysis of 7.





16

Experimental Section

Benzene-*o*-disulfinic Anhydride (7). Sodium sulfite (1.0 g, 7.9 mmol) and sodium carbonate (1.3 g, 12.3 mmol) were dissolved in 10 mL of water and the solution was heated to 40 °C. Solid benzene-*o*-disulfonyl chloride[12] (1.0 g, 3.7 mmol) was gradually added to this solution with continuous stirring. After the addition was complete the mixture was kept at 40 °C for an additional hour; the temperature was then raised to 70–80 °C and stirring continued for 1.5 h. The clear solution was cooled to ice-bath temperature and acidified to pH <1 by the addition of cold, 30% sulfuric acid, and a stream of nitrogen was passed through the solution to remove the sulfur dioxide liberated. A solution of 0.25 g (3.6 mmol) of sodium nitrite dissolved in the minimum amount of water was then added. and the precipitate that formed was filtered off, washed with a small amount of ice–water, and dried at room temperature under vacuum, giving 0.80 g (92%) of benzene-*o*-disulfonhydroxyimide, mp 88 °C (lit.[5] mp 90–91 °C).

Benzene-*o*-disulfonhydroxyimide (0.47 g, 2 mmol) was gradually added with good stirring to 4 mL of 3 N ammonia. Once the addition was complete the solution was allowed to stir for 10 min until gas evolution had completely ceased. The solution was then cooled to 0 °C and acidified to pH <1 with cold 30% sulfuric acid. The acidified solution was extracted with nine 10-mL portions of methylene chloride. The methylene chloride extracts were dried ($MgSO_4$) and the solvent was removed under reduced pressure. The residue was dissolved in 5 mL of methylene chloride at room temperature; addition of 10 mL of hexane led to the crystallization of benzene-*o*-disulfinic anhydride (7), 0.12 g (32%), mp 122–123 °C. The spectral properties of 7 were as follows: IR (KBr) 3088 (w), 1575 (w), 1440 (m), 1145 (s), 1120 cm^{-1} (s); NMR ($CDCl_3$) δ 7.88–8.10, (AA′BB′ pattern); UV (anhydrous dioxane) λ_{max} 272 nm (ϵ 1600). Upon addition of 0.35 mL of 0.1 M aqueous perchloric acid the ultraviolet spectrum of the solution underwent a rapid, but not instantaneous, change to one where the new λ_{max} was at 250 nm.

Although the melting point of the material isolated in the present study (122–123 °C) is slightly higher than that reported by Hendrickson et al.[4] (113–115 °C), the similarity in the infrared spectra indicates that the compound isolated here and that obtained by Hendrickson and co-workers[4] are the same.

Preparation and Purification of Other Materials. Phenyl benzenesulfinyl sulfone[13] was prepared from benzenesulfinic acid by using the procedure for synthesis of sulfinyl sulfones from sulfinic acids developed by Moffatt and Lerch.[14] Dioxane was purified as described by Fieser and Fieser[15] and stored frozen at −20 °C to prevent the formation of peroxides prior to use. Doubly distilled water was used in all kinetic experiments. Perchloric acid and lithium perchlorate were analytical reagent grade.

Procedure for Kinetic Runs. In the runs with 7 a solution of the sulfinic anhydride (5×10^{-4} M) in anhydrous dioxane was placed in one of the reservoir syringes of a Durrum-Gibson Model D-110 stopped-flow spectrophotometer. A solution containing the desired concentrations of perchloric acid (and lithium perchlorate, where used) in 80% dioxane–20% water was placed in the other reservoir syringe. Upon mixing, the course of the reaction was followed by monitoring the increase in the absorbance of the solution at 250 nm on the storage oscilloscope.

In the runs with 9 the rates were enough slower that conventional spectrophotometry could be used: 3.5 mL of a 90% dioxane–10% water solution containing the desired concentration of perchloric acid was placed in a 1-cm cell in the thermostatted cell compartment of a Cary Model 17 spectrophotometer and the reaction was initiated by the addition and rapid mixing of 35 μL of a solution of 9 (10^{-2} M) in anhydrous dioxane. The decrease in the absorbance of the solution with time at 300 nm was then followed.

Acknowledgment. Support of this research by the Robert A. Welch Foundation (Grant D-650) is gratefully acknowledged.

Registry No. 6, 21691-14-5; 7, 77450-61-4; 9, 784-81-6; 10, 77450-62-5.

---

[11] The σ value for *p*-MeS(O) is +0.5; also an aromatic ring itself is electron-withdrawing compared to an alkyl group.

[12] Parham, W. E.; Roder, T. M.; Hasek, W. R. *J. Am. Chem. Soc.* **1953**, *75*, 1647.

[13] Kice, J. L.; Pawlowski, N. E. *J. Am. Chem. Soc.* **1964**, *86*, 4898.

[14] Moffatt, J. G.; Lerch, U. *J. Org. Chem.* **1971**, *36*, 3686.

[15] Fieser, L. F.; Fieser, M. "Reagents for Organic Synthesis"; Wiley: New York, 1967; Vol. 1, p 333.