Registry No. 1, 16653-41-1; 6, 104034-85-7; 6 ($R^3 = R^4 = H$), 104034-89-1; 7, 104034-86-8; 8, 104034-87-9; 8 ($R^3 = R^4 = H$), 104034-90-4; 9, 104034-91-5; **10a**, 2530-07-6; **10b**, 104034-96-0; **10c**, 104034-97-1; **11b**, 104034-98-2; **11c**, 104034-99-3; 13, 104035-00-9; 14, 104051-49-2; 3-*O*-acetyl-26-*O*-[$^2H_3$]acetyl-5α-22ξ-[5,6,22,23-$^2H_4$]furostan, 104034-88-0; tigogenin-d_3 (isomer 1), 104034-92-6; tigogenin-d_3 (isomer 2), 104034-93-7; neotigogenin-d_3 (isomer 1), 104034-94-8; neotigogenin-d_3 (isomer 2), 104034-95-9.

A Study of the Hydrolysis of Methoxysilanes in a Two-Phase System

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The hydrolysis rates of $(CH_3)_{4-x}Si(OCH_3)_x$, $x = 1$–3, in a two-phase system composed of a hydrocarbon layer and an aqueous buffer phase were found to be first order in substrate and specific acid or base catalyst for pH 4–11. In addition, evidence for catalysis by bicarbonate was noted. In contrast to what one would predict on the basis of literature precedence, under acidic conditions, the following reactivity was found: $(CH_3)_3SiOCH_3 > (CH_3)_2Si(OCH_3)_2 > CH_3Si(OCH_3)_3$, whereas the order was reversed with basic catalysis. Activation parameters are consistent with an associative mechanism with either acid or base catalysts. This study, coupled with results in other solvents, suggests the intermediacy of a five-coordinate silicon structure which in acidic media has both entering and leaving groups on the same face of the alkoxysilane. The transition state or intermediate in the base-catalyzed process is proposed to be analogous to that of the S_N2 reaction of carbon systems. Qualitatively, aqueous solutions of hydrolyzed trifunctional silicon species were found to be the most stable, as characterized by the absence of gel formation, at pH values between 3 and 8. Outside of this range condensation and/or phase separation was noted.

Solvolysis studies of the Si–O bond are amply represented in the literature.[1,2] The compounds investigated have primarily either been aryl/(aryloxy)silanes that have chromophores which facilitate the monitoring of the reaction by UV–vis spectroscopy or carboxysilanes which when hydrolyzed yield an acid that can be measured by simple titration. A systematic study of the hydrolysis of a series of alkylalkoxysilanes has not been published, nor is it easy to compare the available examples as different solvents and various catalysts have been used. The following report is a description of the specific acid/base-catalyzed hydrolysis chemistry of the series $R_{4-x}Si(OR')_x$ where $x = 1$–3 and $R = R' = CH_3$. These compounds, particularly dimethyldimethoxysilane (DMDM) and methyltrimethoxysilane (MTMS), are important to the silicone industry, and their hydrolysis is often viewed as the initial reaction in the so-called "cure"—the process that produces polysiloxane elastomers and resins.[3]

Results and Discussion

Initial studies of proton-catalyzed hydrolysis of MTMS were carried out in an acetone/water solvent mixture. From these experiments, in which water was in relatively large excess with respect to MTMS, were determined first-order dependence for substrate[4] and catalyst, *p*-toluenesulfonic acid, over a range of acid concentrations—10^{-3} to 10^{-5} M. In the study the disappearance of MTMS was monitored by either ^1H NMR spectroscopy (determination of MTMS order) or by GC analysis (determination of catalyst order) in which the method of initial rates was used—the reaction being monitored to only 20–30% completion. Although estimates for activation parameters ($\Delta H^{\ddagger} = 7.1$ kcal mol^{-1}, $\Delta S^{\ddagger} = -30$ eu) were obtained over a 35° range and were consistent with the report of an analogous compound,[2a] several limitations—such as the sampling rate being determined by GC cycle times,[5] the acidity of catalysts in these unbuffered, mixed solvent systems being difficult to measure,[6] and the possibility of competitive processes involving hydrolysis products—led the investigation down a different avenue as described below.

The MTMS–water system exists as two phases if agitation is minimized, and the location of the interface changes with time as the MTMS diffuses into the aqueous layer. A plot of the volume change vs. time is linear ($r = 0.996$ for the following expression: $\Delta V(\text{mL}) = (2.62 \times 10^{-4}\ \text{mL s}^{-1})\ t(\text{s}) + 7.9 \times 10^{-2}$ mL, at pH 5.0 and where $\mu = 0.01$ M), suggesting hydrolysis and/or diffusion is occuring at some constant rate in accord with the concentration of MTMS being constant at the interface (see supplementary material). By using aqueous buffer solutions it was also determined that there existed a pH dependence for the process—qualitatively, the rate was slowest at a pH close to 7.

In light of these observations, a series of reactions were examined in which methylalkoxysilanes dissolved in pentane were mixed with aqueous buffer solutions, and the disappearance of starting material in the pentane phase was monitored. A sequence (Scheme I) in which the reaction is presumed to take place in the aqueous layer is supported by the absence of significant hydrolysis when

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(1) For a review, see: Prince, R. H. *M. T. P. International Reviews of Science Inorganic Series One*, Tobe, M. L., Ed.; Butterworths: London, 1972; Vol. 9.

(2) (a) Pratt, R. F.; McNeil, K. J.; DiCaprio, J. A.; Walsh, D. A. *J. Am. Chem. Soc.* **1980**, *102*, 1859 and references cited therein. (b) Boe, B. *J. Organomet. Chem.* **1972**, *43*, 275 and references cited therein.

(3) Noll, W. *Chemistry and Technology of Silicones*; Academic Press: New York, 1968.

(4) White, M. A.; Smith, K. A., unpublished results.

(5) Attempts to quench the reaction by the addition of base or by cooling the samples were unsuccessful.

(6) Paul, M. A.; Long, F. A. *Chem. Rev.* **1957**, *57*, 1–45. Though not a recent review of acidity functions, this paper gives an excellent treatment of the problem of determining acidity in mixed solvents and when weak acids are employed.

0022-3263/86/1951-3827$01.50/0

Table I. Catalyst Dependence for Hydrolysis of Methylmethoxysilanes: |Slope| of log k_{obsd} vs. pH

compound	acid	r	base	r
MTMS	0.98	0.983	1.06	0.996
DMDM	1.04	0.999	0.98	0.998
$(CH_3)_3SiOCH_3$	1.00	0.995	0.99	0.998

the two phases are separated.[10] This behavior allows one to examine relatively rapid processes by GC, limited only by sampling rate.

Scheme I

$$\text{Si(hydrocarbon)} \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} \text{Si(aqueous)} \qquad 1/K_1$$

$$\text{Si(aqueous)} + H^+ \underset{k_{-2}}{\overset{k_2}{\rightleftharpoons}} \text{Si(aqueous)}\cdot H^+ \qquad K_2$$

$$\text{Si(aqueous)}\cdot H^+ + H_2O \xrightarrow{k_{H^+}} \text{Si(OH)} + CH_3OH$$

$$\text{Si(aqueous)} + {}^-OH \xrightarrow{k_{^-OH}} \text{Si(OH)} + CH_3O^-$$

A solution (eq 1) to the equations of Scheme I is obtained by assuming a fast equilibrium for the protonation

$$\frac{-d[\text{Si(hyd)}]}{dt} = \frac{[\text{Si(hyd)}]}{K_1}\left(K_2k_{H^+}[H^+][H_2O] + \frac{K_wk_{^-OH}}{[H^+]}\right) \tag{1}$$

step, as well as equilibrium between the phases. From this treatment, one anticipates the observed rates to be dependent on the stirring rate unless efficient mixing is achieved. For an unstirred solution of MTMS in pentane over a layer of aqueous buffer (pH 4.1), $k_{obs} = 2.4 \times 10^{-5}$ s^{-1}, while for the same reaction stirred magnetically, $k_{obs} = 1.2 \times 10^{-3}$ s^{-1}. Finally, with rapid mechanical stirring, a maximum ($k_{obs} = 3.0 \times 10^{-3}$ s^{-1}) is observed. With good mixing standard deviations for repeated measurements were 5–7%; although larger values (10%) were recorded for the most rapid runs (supplementary material). From the excellent correlations and good reproducibility, one can suppose that within the pH range, 4–11, one is not diffusion-limited. Further evidence for this conclusion is that under acidic conditions the observed order of reactivity is the reverse of that expected on the basis of solubility (vide infra).

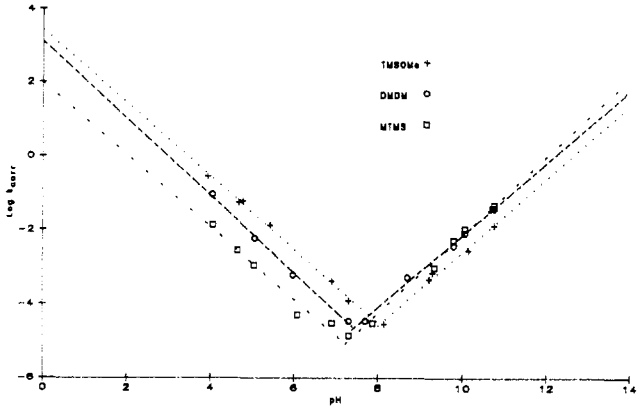

Figure 1. Reaction rate as a function of pH. The linear functions for rates not corrected for K_1 intercept the y axis at 1.32 (MTMS), 2.18 (DMDM), and 2.06 ($(CH_3)_3SiOCH_3$).

Table II. Second-Order Rate Constants for Hydrolysis of Methylmethoxysilanes at 0 °C[a]

	k_{H^+}',[b]	$k_{^-OH}$, M^{-1} s^{-1}
MTMS	98 ± 10	126 ± 15
DMDM	1270 ± 100	58 ± 6
$(CH_3)_3SiOCH_3$	2660 ± 500	24 ± 5
$C_6H_5Si(OCH_2CH_2OCH_3)_3$[c]	47	165

[a] Neither acid or base constants have been corrected for statistical factors, ref 9. [b] $k_{H^+}' = K_2k_{H^+}$. [c] At 30 °C, ref 2a.

Plots of ln [substrate] vs. time show good linearity over several half-lives. Runs for pH 4–6 or pH 8–11 were taken to at least 3 half-lives, and in the case of DMDM and MTMS, good correlation ($r > 0.99$) was found in either acidic or basic media. Significant condensation to gels was observed for MTMS at pH values >8, yet the correlation was still acceptable. For trimethylmethoxysilane ($TMSOCH_3$), the correlation with the first-order model was good on the strong acid or base side, but poor to fair (>0.98) at near neutral conditions. Trimethylsilanol produced in the hydrolysis was partitioned between the two phases as evidenced in the GC trace, yet no condensation product, hexamethyldisiloxane, was observed in any of the runs. Within the pH range 6–8, hydrolysis of all substrates was monitored for much less conversion, as little as 5%, for the extremely slow reactions ($t_{1/2} > 10$ h). Thus, uncertainties in this range are doubtless greater than for the more conveniently paced processes.

A plot of log k vs. pH (Figure 1) displays first-order dependence for both branches (acid and base) of each substrate (Table I). No uncatalyzed hydrolysis was observed for any of the compounds.

The reactions of Scheme I suggest an important parameter in the determination of the actual reactivity of the methoxysilanes to hydrolysis—namely, the substrate's distribution coefficient. By making the assumption that the distribution coefficient does not change markedly with pH, one can measure its value under those conditions where the observed rate is close to its minimum and correct over the entire pH range. Solubility equilibria—assumed to be as follows

$$K_1 = \frac{[\text{Si(hydrocarbon)}]}{[\text{Si(aqueous)}]}$$

were determined at pH values of 6.9 to 7.3 and $\mu = 0.1$ M. For pentane at 0 °C, log K_1 values were as follows: MTMS, 0.67 ± 0.05; DMDM, 0.92 ± 0.05; $(CH_3)_3SiOCH_3$, 1.36 ± 0.21. Correcting for these constants reveals the true differences in relative reactivity (Figure 1 and Table II). For the acidic side one sees considerable discrimination among the substrates—k_{H^+}' ranging over 30-fold, while this feature is attenuated with base catalysis. With alkaline conditions methyltrimethoxysilane was found to hydrolyze approximately five times faster than the monomethoxy substrate.

Additional data suggesting the effect of solubility on rate is the decrease in k_{obs} with an increase in ionic strength

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[7] Structure 1 in the limit would be a square-pyramidal molecule with both leaving and entering groups basal, although, certainly a lower symmetry, C_s, structure could also pertain. In any case, an axial position is assumed by both entering and leaving groups as the five-coordinate intermediate structure develops.

[8] Boe, B. *J. Organomet. Chem.* **1973**, *57*, 255–260, and references cited therein.

[9] For a comparison of these rate constants, statistical corrections made to the acid side—through K_2—increase k_{H^+}' of MTMS and DMDM by factors of 3 and 2, respectively; on the base side, $k_{^-OH}$ of MTMS and DMDM are reduced by factors of 3 and 2, respectively. Likewise, small corrections involving these changed rate constants are manifested in the entropies of activation for MTMS and DMDM.

[10] As noted by a referee, these observations are equally consistent with the antipodal explanation that hydrolysis occurs in the organic layer and that once the sample is removed from contact with the aqueous layer, the remaining water is consumed and the reaction apparently stops. If this explanation is correct, then based on the rather low solubility of water in pentane, the hydrolysis reaction is *indeed* a rapid process.

Table III. Activation Parameters for Hydrolysis Reaction of Methylmethoxysilanes

	acid		base	
	$\Delta H^{\ddagger}$, kcal mol^{-1}	$\Delta S^{\ddagger}$, eu	$\Delta H^{\ddagger}$, kcal mol^{-1}	$\Delta S^{\ddagger}$, eu
MTMS	6.4 ± 0.5	-34 ± 4	10.3 ± 0.5	-14 ± 3
DMDM	8.2 ± 0.5	-25 ± 3	12.0 ± 0.5	-8 ± 2
$(CH_3)_3SiOCH_3$	[a]		8.1 ± 0.5	-16 ± 4[b]

[a] Not determined. [b] These values are apparent activation parameters. See ref 11.

(supplementary material). The net effect is to "salt out" the reactants and, thus, slow the observed process. Interestingly, the maximum rate for MTMS hydrolysis appears to be reached at $\mu = 0.10$ M with k_{obs} decreasing slightly for $\mu < 0.10$, although the data is not accurate enough to be certain of this. All rate constants are reported for $\mu = 0.10$.

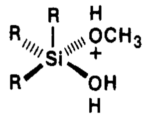
The relative reactivity as a function of pH is predictable on the basis of the relative electronegativity of oxygen compared to carbon. That is, the basicity of an oxygen in MTMS will be less than that of the oxygen of $(CH_3)_3SiOCH_3$ and, therefore, less easily protonated. Likewise, the silicon of MTMS will be more positive than that of $(CH_3)_3SiOCH_3$ and, thus, more susceptible to attack by hydroxide. These simple arguments are useful, but possibly misleading in that additional data on related systems suggest a somewhat more complicated situation.

From his work on the acid-catalyzed alcoholysis of di- and trimethylalkoxysilanes, Boe found that, in benzene, the dialkoxysilanes were much more reactive than the corresponding trimethylalkoxysilanes.[2b] In our labs, it has been observed that, with acid catalysis, methyltrialkoxysilanes hydrolyze faster than dimethyldialkoxysilanes in acetone with 5% water.[4] These data indicate an inverse order to that measured in the two-phase experiment.

To investigate this issue activation parameters were determined for some of the substrates. With decane as the hydrocarbon phase, solubility equilibria were determined as a function of temperature. The measurements were fit to linear plots from 0 to 60 °C and had the following parameters for the functional form, $\log K_1 = b - (a/\text{temp (K)})$: MTMS – $a = 3.57 \times 10^2$ K, $b = 1.87$; DMDM – $a = 7.70 \times 10^2$ K, $b = 3.50$.[11] These data along with the standard measurements of k_{obs} with temperature yielded estimates for the desired quantities (Table III).

Activation parameters for the acid-catalyzed hydrolysis are in general agreement with reports of related compounds or processes[1,2] and with those determined for MTMS in acetone/water (see above). The values indicate an associative reaction. Interestingly, the apparent activation energy for DMDM hydrolysis is greater than that for MTMS—the difference in reactivity arising in the entropy term of the Arrhenius expression.

A rationalization of the above data and that of others may be found by considering the structure of the five-coordinate intermediate to be as depicted by 1,[7] rather than the typically supposed arrangement 2. Intermediate 1, which could arise either by direct addition of water to the same side of the molecule as the protonated alkoxy leaving group or via 2 by pseudorotation, should in the case where all R = alkyl groups have a larger dipole difference between starting material and intermediate than the same structure with either one or two of the alkyl ligands replaced by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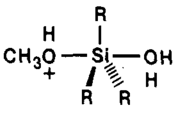
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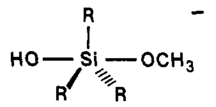
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alkoxy substituents. The relatively more polar form is stabilized in water, while this effect is absent in less polar media. Additional data for this hypothesis is the pronounced difference (approximately 5 kcal mol^{-1}) in E_a for n-propanolysis vs. 2-propanolysis of 2,2-dimethyl-2-sila-1,3-dioxolane.[2b] Since the nucleophile and leaving group in a structure such as 1 are positioned for greater interaction, what would in 2 be a relatively minor perturbation (i.e., branching in the alkoxy) is accentuated in 1.

Apparent activation energies for the base-catalyzed reaction are in general larger than those with acid catalysis throughout the series of compounds. The activation entropies are also larger than those on the acid side, yet are still consistent with an associative process. It should be noted, however, that catalysis by bicarbonate is present in these measurements; therefore, any conclusions one might draw are fraught with uncertainty. Catalysis by bicarbonate was found in MTMS and DMDM (supplementary material), while phosphate catalysis was not discernible against the accummulated error of the method. General base catalysis has previously been demonstrated in alkoxysilane solvolysis.[1,2,8] No acetate catalysis was observed, nor was acetic acid an effective catalyst as determined by this technique. The latter conclusion was also deduced from experiments in the acetone/water system with acetic acid as catalyst.

Finally, attempts to examine the hydrolysis of tetramethoxysilane proved impossible under basic conditions. The plots of concentration vs. time indicated the reaction was not first order in substrate, due apparently to facile condensation reactions. Such was not the case in acid. In this medium, $(CH_3O)_4Si$ apparently hydrolyzed the slowest of all the compounds investigated. More detailed study of this system was not undertaken.

Conclusions

The two-phase experiment can yield data for systems which are not readily analyzed by more conventional procedures. The errors in the method arise primarily from diffusion considerations and changing solubility equilibria. In the systems investigated these problems were generally tractable since the hydrolysis products were much more soluble in the aqueous layer than the organic, and the reactions were not too fast.

The relative reactivity of the methoxysilanes, though predictable on the basis of acidity and electronegativity arguments, when coupled with solvent effects suggests two different structures for the five-coordinate intermediate of the acid- and base-catalyzed processes. The proposed structure for the acid-catalyzed reaction approximates that depicted by 1, while the intermediate (or transition state) of basic hydrolysis is shown as 3.

Experimental Section

Dimethyldimethoxysilane and methyltrimethoxysilane were obtained from Silar and distilled from sodium and sodium methoxide. Trimethylmethoxysilane was prepared from tri-

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(11) Measurement of $K_1(T)$ for $(CH_3)_3SiOCH_3$ was sufficiently irreproducible for temperatures above 35 °C that this dependence was not reported or included in the determination of activation parameters for $(CH_3)_3SiOCH_3$ (see Table III).

methylchlorosilane and methanol and distilled from sodium methoxide. Tetramethoxysilane (Silar) was used as received. Aqueous buffers were prepared with distilled water which was first passed through an ion-exchange resin and then through a series of Millipore filters. The pH of the prepared buffers was measured with a Leeds-Northrup pH meter equipped with a Beckman combination glass electrode. Calibration was done by bracketting each value with standard solutions. Both sodium and potassium salts were used in the experiments and, within experimental error, no difference was detected. Ionic strength of the solutions was obtained either by the buffer system itself or, in the experiments at lower buffer concentrations, was adjusted to the desired value by the addition of sodium perchlorate or, in a few cases, potassium chloride. No difference was noted by the substitution of perchlorate with chloride.

Analysis of the organic layer was performed on a Shimadzu GC-9A with C-R1B data analyzer. Noting the difficulties in reproducibility of FID methods with silicon compounds, we employed TC detection in conjunction with either a 6 ft 3% OV-101 or a 6 ft Porapak Q column. The latter facilitated monitoring water, methanol, and silicon species. Internal standards were decane (Aldrich, 99%) or octane (Aldrich, 99%) with which response factors were determined.

Temperatures were maintained to ±0.2° in a circulating water bath (27 L volume) equipped with a B. Braun Thermomix heater/circulator. Ice-water baths were used for 0 °C temperature.

Two-Phase Experiment. A 100.0-mL solution of 0.10–0.15 M alkoxysilane and 40 mM octane in pentane and a 100.0-mL buffer solution were separately equilibrated to the desired temperature. The solutions were combined in a round-bottomed flask equipped with a mechanical stirrer and septum for removing sample aliquots. The sealed system was mixed vigorously for timed intervals with sampling periods interpersed. During sampling the mixing was stopped and the phases were allowed to separate. Typically this process took 5–10 s. In experiments with MTMS or $(CH_3O)_4Si$ in base, persistent emulsions formed and aliquots were removed without complete segregation of the phases. The samples were analyzed by GC analysis.

Registry No. $(CH_3)_3Si(OCH_3)$, 1825-61-2; $(CH_3)_2Si(OCH_3)_2$, 1112-39-6; $CH_3Si(OCH_3)_3$, 1185-55-3.

Supplementary Material Available: Table of observed rates as a function of pH and substrate, figure of volume change vs. time for MTMS, and figure of disappearance of MTMS vs. time (5 pages). Ordering information is given on any current masthead page.

Mechanistic Aspects and Profiles of the Double Elimination Reaction of β-Substituted Sulfones

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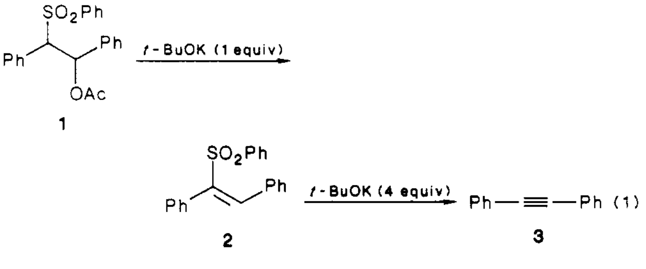
The double elimination reaction of β-acetoxy or β-alkoxy sulfones was investigated in detail by employing some representative reactions. Successful isolation of reaction intermediates revealed the reaction path: the first step is elimination of the acetoxy or alkoxy group to afford a vinyl sulfone. The subsequent elimination of a phenylsulfonyl group from the vinyl sulfone gives acetylenes, while polyenes are formed in cases where isomerization of the vinyl sulfone to an allyl sulfone is possible prior to the second elimination. Besides these mechanistic considerations, general features are discussed in order to make clear the scope and limitations of this synthetically useful reaction.

Previously, we reported a novel double elimination reaction of β-acetoxy or β-alkoxy sulfones which afforded a convenient method for synthesizing enynes and polyenes (Scheme I).[1] Acetylenes are formed when R_2 are phenyl, phenylethynyl, furyl, thienyl, and β-(triorganosilyl)vinyl or -ethynyl groups. By contrast, other (substituted) alkyl and alkenyl aldehydes provide polyene compounds. The utility of this reaction was exemplified by the successful synthesis of *d,l*-muscone and methyl retinoate. More recently, we have demonstrated that this method could be applied to the novel synthesis of vitamin A[2] and (2*E*,4*E*)-dienamides and -dienoates[3] of greater than 90% stereochemical purity. It seems of importance, therefore, to draw out the profile of this reaction. Of particular interest is to disclose what factors determine the reaction path to give either the acetylenic or polyenic compounds. In the hope of obtaining better understanding of these problems, we were attracted to an investigation of some representative reactions in detail. Actually, we have succeeded in isolating the reaction intermediates. In this paper, we describe the mechanistic aspects as well as make some general remarks on the double elimination reaction to make clear the scope and limitations.

Results

Acetylene Formation. When the β-acetoxy sulfone **1** was treated with 1 equiv of *t*-BuOK in *t*-BuOH at room temperature, the vinyl sulfone **2** was formed immediately in 94% yield (eq 1). This is quite reasonable since the hydrogen α to the phenylsulfonyl group is more acidic than the one α to the acetoxy group. The analogous vinyl sulfone formation from the corresponding tosylates and



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[1] Mandai, T.; Yanagi, T.; Araki, K.; Morisaki, Y.; Kawada, M.; Otera, J. *J. Am. Chem. Soc.* **1984**, *106*, 3670.

[2] (a) Otera, J.; Misawa, H.; Mandai, T.; Onishi, T.; Suzuki, S.; Fujita, Y. *Chem lett.* **1985**, 1883. (b) Otera, J.; Misawa, H.; Onishi, T.; Suzuki, S.; Fujita, Y. *J. Org. Chem.*, following paper in this issue.

[3] Mandai, T.; Moriyama, T.; Tsujimoto, K.; Kawada, M.; Otera, J. *Tetrahedron Lett.* **1986**, *27*, 603.

0022-3263/86/1951-3830$01.50/0