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Kinetics and Thermodynamics of Phenolate Silylation and Alkylation

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Abstract: A kinetic and thermodynamic investigation of the silylation reactions of alkali phenolates with several trisubstituted silyl chlorides was performed in tetrahydrofuran, acetonitrile, and dimethyl sulfoxide. Heats and rates of reaction were determined by titration calorimetry and stopped-flow techniques and found to be strongly dependent upon solvent polarity, ion pairing of the alkali phenolate, and steric and electronic influences. Reaction rates were considerably faster in acetonitrile and dimethyl sulfoxide than in tetrahydrofuran where ion-pairing effects were significant. Rates were accelerated by the addition of crown ethers. In most cases, potassium phenolates were more reactive than the corresponding sodium phenolates. The order of reactivity observed for the silyl chlorides increased as follows: $i\text{-Pr}_3 < t\text{-BuPh}_2 < \text{Me}_3 < \text{Ph}_2\text{Me}$. Substitution with *tert*-butyl functionalities in the ortho position of the phenolate ring retarded reactivity through steric crowding of the reaction site. Resonance and/or inductive effects of substituents increased the reactivities of phenolate anions in the following order: $2,4,6\text{-(NO}_2)_3 < 2,4\text{-(NO}_2)_2 < 2,6\text{-(}t\text{-Bu)}_2\text{-4-NO}_2 < 4\text{-NO}_2 < 4\text{-CN} < 4\text{-Cl} < 4\text{-F} < 4\text{-H} < 4\text{-}t\text{-Bu} < 4\text{-OMe}$. Linear relationships were found between the heats of reaction, the pK_a's of the parent phenols, and the appropriate Hammett σ constants. Hammett plots for the silylation reactions of substituted potassium phenolates with triphenylsilyl fluoride gave values of $\rho = 1.24$ and 2.40 for reaction in acetonitrile and dimethyl sulfoxide, respectively.

Base-promoted reactions of carbonyl compounds are the most commonly employed processes in modern synthetic organic chemistry. However, because of their complexity, these systems have been the subjects of relatively few systematic physical organic studies.[1-5] Reactions of alkali enolates are currently carried out at low temperatures in solvents of low dielectric constant and low basicity where ion pairing and higher degrees of aggregation occur and have an important bearing on the reaction mechanisms and the stereochemistry of the products.[3-13]

In recent years, many of the difficulties in handling unstable enolates have been circumvented by the use of silyl enol ethers as relatively stable synthons whose regioisomers can be isolated and separated prior to reaction. Furthermore, House et al.[14] and Stork and Hudrlik[15a,b] have shown how to generate selectively the kinetic and thermodynamic trimethylsilyl enol ether isomers of ketones. This approach has suddenly brought silyl ethers from a relatively obscure position to one of great interest and importance to contemporary organic chemistry.

The purpose of this article is to report our kinetic and thermodynamic comparisons for the silylation of a series of alkali phenolates and a β-diketo enolate in several organic media. We will also compare silylation reactions to methylation of the same phenolates under comparable conditions. Relatively few rate studies are available for the silylation of enolates, phenolates, or alkoxides primarily because of the very high reactivity of silyl chlorides, the reagents used most commonly for the preparation of silyl enol ethers. However, Stork and Hudrlik[15b] did note that the rate of *cleavage* of silyl enol ethers was strongly dependent on the type of organometallic reagent used, the solvent, and the alkyl substituents on the silyl group. Corriu and Guerin[16] observed that the stereochemistry and mechanism of silylation for substituted phenolates with asymmetric silylating agents was influenced by the electronic environment of the substituent.

Several kinetic studies of silyl compounds emphasize the importance of steric effects.[17-19] The relative solvolysis rates of triphenylsilyl chloride and triisopropylsilyl chloride fall in the following order of the solvent: water > methanol > 1-propanol > 1-hexanol > 2-propanol. Akerman[17] obtained solvolysis rates for several trialkyl-substituted phenoxysilanes in aqueous ethanol. Reactivity decreased with increasing steric interactions in the following order: $\text{Me}_3 > \text{Et}_3 > n\text{-Pr}_3 > n\text{-Bu}_3 > n\text{-Pen}_3 > t\text{-BuMe}_2$. Silylation of *p*-nitrophenol with several bis(trialkylsilyl)acetamides in dioxane gave the following rate order: $\text{Me}_3 > \text{Me}_2\text{Et} > \text{Et}_3$. Electronic factors may also influence displacement reactions at silicon.[16,20,21] Just as displacement at carbon may occur through a variety of mechanisms ($\text{S}_\text{N}1$, $\text{S}_\text{N}2$, SET), displacement reactions involving silyl halides also proceed through several mechanisms.[22-24]

In view of all the complications described above, we have chosen to study the reactions of silyl halides with alkali phenolates, because of their relatively low reactivity compared to aliphatic enolates and to ensure attack at oxygen rather than the intrusion of ambient reactivity.[5,25] Furthermore, phenolates provide the opportunity for modulating electronic effects through substitution at the para position or steric effects by substitution at the 2- and

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6-positions. Even with these simplifications the comparison of alkali phenolates in nonhydroxylic solvents (the only kind appropriate for silylation studies) must be corrected, as we have done previously,[26a] through the Acree equation provided the degree of dissociation of the ion aggregates has been determined.

In this report we will present comparison of second-order silylation rates in tetrahydrofuran and acetonitrile at 25 °C for the reaction of several silyl chlorides with lithium, sodium, and potassium phenolates whose association constants have been determined by conductance measurements under the same conditions as the rate studies. In addition to variation of ion association through change of solvent, the polybasic ligands, i.e., crown ethers and cryptands, were used to produce "naked" phenolate anions for making conductance and rate comparisons. Unfortunately, the silyl chlorides were so reactive and their rates of reaction covered such a wide range that we were unable to compare them without special manipulation of their rates to a measurable range by deactivation of the phenolate system by progressive nitration of the ring and the introduction of bulky groups at the 2- and 6-positions. However, a more extended rate comparison of a series of para-substituted phenolate ions was feasible using the much less reactive triphenylsilyl fluoride.

Reaction calorimetry provided the means for comparing the heats of formation of the Si–O bonds for the entire series of substituted phenolates and silyl chlorides and fluoride. The heats of formation of these bonds can be converted to the heats of Si–O heterolysis simply by changing the sign. This provides, to the best of our knowledge, the first quantitative comparisons of heterolysis energies for a series of silicon–oxygen bonds in solution.

Experimental Section

The compounds used in this study were available commercially or prepared via standard literature procedures. In general, compound purity was assessed by proton NMR, boiling point, and melting point (Thomas-Hoover capillary apparatus). All compounds were stored under argon or in a vacuum desiccator to prevent decomposition. Solvents were prepared as described below and were either distilled just before use or stored under an argon atmosphere in a dispensing buret. All glassware was dried at 125 °C for at least 4 h prior to use. For further details see ref 27.

Dimethyl sulfoxide, DMSO, was purified by the method of Arnett et al.[28] with minor modifications.

Dimethyl-d_6 sulfoxide (99.9 atom % D) was obtained from MSD Isotopes and was used directly from the ampule.

Acetonitrile (Mallinckrodt) was predried over 3-Å molecular sieves and was distilled subsequently from phosphorus pentoxide (5 gm/L) under argon. Only the center fraction (~75%) was retained from the distillation. This solvent was used only if the water content was less than 25 ppm by Karl Fischer titration. The acetonitrile was stored under argon in a solvent bottle equipped with a dispensing buret.

Acetonitrile-d_3 (99 atom % D) was obtained from Aldrich and used as received.

Tetrahydrofuran (Fisher, ACS grade), THF, was distilled just prior to use from sodium and benzophenone if the blue color of the benzophenone ketyl was persistent.

Silylating agents were used as received (from Aldrich or Petrarch) after careful examination by 300-MHz proton NMR for purity and hydrolysis. Triphenylsilyl fluoride was recrystallized from benzene prior to use.

The phenols were used as received or were recrystallized from petroleum ether or methanol. Purity was checked by proton NMR and/or melting point.

2,6-Di-*tert*-butyl-4-nitrophenol was prepared by the procedure of Stroh et al.[29] with minor modifications. 2,6-Di-*tert*-butylphenol (103 gm, 0.5 mol) was dissolved in hexane with heating to 50 °C. A 95-mL aliquot of 30% w/w nitric acid/water was added to the hexane solution over the course of 1 h. The reaction temperature was maintained between 50 and 55 °C while the reaction ran overnight. The yellow precipitate, 2,6-di-*tert*-butyl-4-nitrophenol, was filtered and washed three times each with hexane followed by water. The yellow solid was dried under vacuum: mp 155–156 °C (lit.[29] mp 156 °C); ^{1}H NMR ($CDCl_3$) δ 8.12 (s, 2 H, meta), 5.93 (s, 1 H, OH), 1.45 (s, 18 H, CH_3). Dibenzoylmethane (1,3-diphenyl-1,3-propanedione; Aldrich) was purified by recrystallization from petroleum ether and multiple recrystallizations from methanol: mp 75–76 °C (lit.[30] mp 75–76 °C); ^{1}H NMR (DMSO-d_6) δ 8.17 (d, 4 H, meta), 7.66 (t, 2 H, para), 7.57 (t, 4 H, ortho), 7.36 (s, 2 H, CH_2).

Alkali phenolates were prepared in the following manner. Methanol was distilled from sodium metal to reduce the water content. Freshly shaved potassium, sodium, or lithium was weighed under dry benzene and brought into reaction with the freshly distilled methanol to generate the alkali methoxide under argon. Approximately a 1% excess of phenol was added and the solution stirred for $^1/_2$ h. A small amount of benzene was also added to azeotrope out traces of water; then the solvent was removed by rotary evaporation. The resulting solid was washed with hexane to remove excess phenol and dried under high vacuum. All alkali phenolates were stored in a vacuum desiccator. ^{1}H NMR spectra of the alkali phenolates and the β-diketonate in DMSO-d_6 or CD_3CN were used as the primary test for purity.

Nuclear Magnetic Resonance. Proton NMR spectra were obtained on an IBM Instrument NR/80 spectrometer operating at 80 MHz or a Varian XL-300 NMR spectrometer operating at 300 MHz. Heteronuclear NMR spectra were obtained either on the Varian XL-300 or a JEOL FX-90Q Fourier transform NMR spectrometer. All chemical shifts are relative to tetramethylsilane.

High-Pressure Liquid Chromatography. In some cases, reaction products were purified by reverse-phase HPLC on a Varian 5000 liquid chromatograph equipped with a variable-wavelength UV detector. Samples were analyzed on an analytical column (Alltech, 5-μm reverse-phase C18 packing, 4.6 mm × 25 cm) or isolated from a preparative column (Rainin Dynamax Modular Macro-HPLC column system, 8-μm reverse-phase C18 packing, 21.4 mm × 25 cm). The solvent system usually employed was a mixture of tetrahydrofuran/water (60/40).

Silylation Products.

Trimethylphenoxysilane: ^{1}H NMR (CD_3CN) δ 7.21 (t, J = 8.0 Hz, 2 H, meta), 6.86 (d, J = 7.6 Hz, 1 H, para), 6.80 (d, J = 8.7 Hz, 2 H, ortho), 0.07 (s, 9 H, CH_3).

Trimethyl(4-cyanophenoxy)silane: ^{1}H NMR (CD_3CN) δ 7.57 (d, J = 8.6 Hz, 2 H, meta), 6.93 (d, J = 8.6 Hz, 2 H, ortho), 0.07 (s, 9 H, CH_3).

Trimethyl(4-fluorophenoxy)silane: ^{1}H NMR (CD_3CN) δ 6.94 (m, J = 8.8, 2.0 Hz, 2 H, meta), 6.78 (m, J = 4.3 Hz, 2 H, ortho), 0.07 (s, 9 H, CH_3).

Trimethyl(4-nitrophenoxy)silane: ^{1}H NMR (CD_3CN) δ 8.12 (d, J = 9.2 Hz, 2 H, meta), 6.94 (d, J = 9.2 Hz, 2 H, ortho), 0.07 (s, 9 H, CH_3).

Trimethyl(2,6-di-*tert*-butyl-4-nitrophenoxy)silane: ^{1}H NMR (CD_3CN) δ 8.10 (s, 2 H, meta), 1.46 (s, 18 H, CH_3), 0.07 (s, 9 H, $Si(CH_3)_3$).

Diphenylmethyl(4-cyanophenoxy)silane: ^{1}H NMR (CD_3CN) δ 7.69 (m, J = 2.3 Hz, 4 H, meta′), 7.57 (d, J = 8.7 Hz, 2 H, meta), 7.52 (d, J = 3.1 Hz, 2 H, para′), 7.42 (d, J = 2.4 Hz, 4 H, ortho′), 6.97 (t, J = 8.7 Hz, 2 H, ortho), 0.62 (s, 3 H, CH_3).

Trimethyl(dibenzoylmethoxy)silane: ^{1}H NMR (CD_3CN) δ 8.08 (d, J = 4.3 Hz, 4 H, meta), 7.63 (t, J = 7.3 Hz, 2 H, para), 7.55 (t, J = 8.1 Hz, 4 H, ortho), 7.08 (s, 1 H, vinyl), 0.95 (s, 9 H, CH_3).

Triisopropyl(dibenzoylmethoxy)silane: ^{1}H NMR (CD_3CN) δ 8.08 (d, J = 7.0 Hz, 4 H, meta), 7.62 (t, J = 7.0 Hz, 2 H, para), 7.54 (t, J = 7.9 Hz, 4 H, ortho), 7.07 (s, 1 H, vinyl), 1.25 (q, J = 6.5 Hz, 3 H, methine), 1.04 (d, J = 3.6 Hz, 18 H, CH_3).

Kinetics. Second-order rate constants from 10^{-3} to 10^1 L·mol^{-1} s^{-1} were obtained on a stopped-flow apparatus designed by Dr. G. W. Schriver.[27]

A Durrum-Gibson stopped-flow kinetics apparatus interfaced with a North Star Model 3820 data system was used to measure second-order reaction rates faster than 10^1 L·mol^{-1} s^{-1}. The sample cell had a volume of 200 μL and an optical path length of 2.0 cm. Accumulation of data points could be obtained at 1-ms intervals.

For relatively slow reactions, with half-lives greater than 15 min, rates were measured on a Varian DMS 100 UV-visible spectrophotometer.

All experiments were run under pseudo-first-order conditions, generally with a 100–1000-fold excess of alkylating agent or silylating agent. A decrease in absorbance of the enolate anion (1–10 mM) was monitored at an appropriate wavelength. Rate constants and standard deviations were determined from six to eight replications of each reaction. The pseudo-first-order rate constant, k_{obsd}, was then converted to the second-order rate constant for the reaction. Reaction rates were consistent from one instrument to another as demonstrated by the reaction of triphenylsilyl fluoride with potassium 4-nitrophenolate in DMSO at 25 °C,

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Table I. Second-Order Rate Constants for the Methylation of Sodium Phenolate (NaOPh) with Methyl Iodide (1.69 M) in Acetonitrile and Tetrahydrofuran at 25 °C with and without 15-Crown-5 as an Addendum

[NaOPh], M × 10^3	[addendum], M × 10^3	k_2, L·mol^{-1} s^{-1}	
		acetonitrile	THF
17.90		$(9.37 \pm 0.44) \times 10^{-3}$	$(1.67 \pm 0.14) \times 10^{-3}$
8.95		$(2.25 \pm 0.06) \times 10^{-2}$	$(1.76 \pm 0.06) \times 10^{-3}$
4.47		$(5.55 \pm 0.12) \times 10^{-2}$	$(4.15 \pm 0.18) \times 10^{-3}$
2.24		$(1.70 \pm 0.07) \times 10^{-1}$	$(5.86 \pm 0.09) \times 10^{-3}$
1.12		$(1.97 \pm 0.08) \times 10^{-1}$	
17.90	17.90	$(7.92 \pm 0.11) \times 10^{-1}$	$(3.01 \pm 0.09) \times 10^{-1}$
8.95	8.95	1.06 ± 0.03	$(2.69 \pm 0.06) \times 10^{-1}$
4.47	4.47	1.22 ± 0.15	$(2.57 \pm 0.10) \times 10^{-1}$
17.90	89.50	$(8.66 \pm 0.18) \times 10^{-1}$	
8.95	44.70	1.06 ± 0.01	
4.47	22.35	1.26 ± 0.06	

Table II. Second-Order Rate Constants for the Silylation of a Series of Para-Substituted Potassium Phenolates (KOPh-X) with Triphenylsilyl Fluorides in Dimethyl Sulfoxide and Acetonitrile at 25 °C

X	k_2, L·mol^{-1} s^{-1}	
	DMSO	acetonitrile
$C(CH_3)_3$	$(1.28 \pm 0.06) \times 10^{3}$	
Cl	$(6.61 \pm 1.04) \times 10^{2}$	10.60 ± 0.43
CF_3	$(3.23 \pm 0.14) \times 10^{2}$	2.91 ± 0.02
C_6H_5	$(2.10 \pm 0.34) \times 10^{2}$	
$COOCH_3$	$(1.72 \pm 0.05) \times 10^{2}$	
CN	7.58 ± 0.38	1.16 ± 0.18
NO_2	$(2.21 \pm 0.07) \times 10^{-1}$	$(7.16 \pm 0.58) \times 10^{-1}$

a reaction whose rates overlap the range of both stopped-flow instruments.

Calorimetry. Heats of reaction were measured by titration calorimetry with a Tronac Model 450 calorimeter operated in the isoperibol mode. In a typical experiment, 50 mL of a 5–10 mM alkali phenolate solution was placed in the reaction vessel and the apparatus was assembled. Approximately 1.5 mL of a solution of silylating agent (0.1–0.2 M) was placed in a calibrated Gilmont syringe from which it was used as the titrant of the phenolate solution in the usual manner.[26b] Three to six titrations were made with each stock solution. In most cases, the reported heats of reaction and standard deviations were obtained from three separate runs.

Conductance. Conductance experiments were performed with a Beckman RC18-A conductivity bridge and a Beckman Jones type conductance cell equipped with two platinum disk electrodes ~2.5 cm in diameter and suspended ~0.2 cm apart. The cell constant was determined to be 5.322×10^{-2} cm^{-1} by using a 0.01 D aqueous KCl solution at 25 °C.

Stock solutions of known alkali phenolate concentrations in tetrahydrofuran were prepared under argon and equilibrated at 25 °C. Sample solutions (1.0×10^{-4}–3.0×10^{-2} M) were prepared by dilution of the stock solution using volumetric flasks. All glassware was rinsed with deionized water and dried at 125 °C overnight prior to use. The conductance cell was then filled to a known level with the sample solution and immersed in a 25 °C constant-temperature water bath. The solution was allowed to equilibrate for at least 30 min prior to each measurement. The conductance cell was rinsed several times with tetrahydrofuran and twice with the next sample prior to refilling the cell.

Results

The aim of this project is to provide quantitative comparisons of the reactivities of various silylating agents with the alkali salts of phenolates and enolates. It was soon obvious from preliminary experiments that a complete kinetic study would be seriously limited due to the very high reactivities of the silyl chlorides which are the most commonly used silylating reagents. Furthermore, as might be expected for a rate study of salts of this type in nonhydroxylic solvents, rates of reaction were highly dependent upon the concentration of the salt, the particular alkali cation employed, the solvent, and the presence of cyclic polybasic ligands such as crown ethers or cryptands. Ion pairing is clearly implied by this behavior. Results will be presented in terms of second-order rate constants as a function of the above variables.

Table III. Second-Order Rate Constants for the Silylation of Sodium 2,4-Dinitrophenolate (NaDNP) with *tert*-Butyldiphenylsilyl Chloride (t-BuPh$_2$SiCl = 0.404 M) and Triisopropylsilyl Chloride (i-Pr$_3$SiCl = 0.409 M) in Tetrahydrofuran at 25 °C with and without 15-Crown-5 as Addendum

[NaDNP], M × 10^3	[addendum], M × 10^3	k_2, L·mol^{-1} s^{-1}	
		t-BuPh$_2$SiCl	i-Pr$_3$SiCl
4.47		$(3.76 \pm 0.09) \times 10^{-1}$	$(1.01 \pm 0.03) \times 10^{-1}$
2.24		$(6.22 \pm 0.14) \times 10^{-1}$	$(2.58 \pm 0.10) \times 10^{-1}$
1.12		$(8.58 \pm 0.18) \times 10^{-1}$	$(4.42 \pm 0.41) \times 10^{-1}$
0.56		$(9.71 \pm 0.36) \times 10^{-1}$	
4.47	4.47	8.85 ± 0.28	1.63 ± 0.01
2.24	2.24	11.59 ± 0.26	2.20 ± 0.04
1.12	1.12	14.32 ± 0.89	3.63 ± 0.05
0.56	0.56		7.03 ± 0.40

Table IV. Second-Order Rate Constants for the Silylation of Sodium 2,6-Di-*tert*-butyl-4-nitrophenolate (NaDBNP) with Triisopropylsilyl Chloride (i-Pr$_3$SiCl = 0.455 M) and *tert*-Butyldiphenylsilyl Chloride (t-BuPh$_2$SiCl = 0.375 M) in acetonitrile at 25 °C

	k_2, L·mol^{-1} s^{-1}	
[NaDBNP], M × 10^3	i-Pr$_3$SiCl	t-BuPh$_2$SiCl
18.05	1.70 ± 0.01	2.29 ± 0.02
9.03	2.05 ± 0.02	3.02 ± 0.04
4.51	2.35 ± 0.03	3.51 ± 0.06
2.26	2.34 ± 0.03	3.59 ± 0.03
1.13	2.53 ± 0.04	3.14 ± 0.03
0.56	2.54 ± 0.11	3.14 ± 0.10

Table V. Second-Order Rate Constants for the Silylation of Sodium Picrate (NaPic) with Diphenylmethylsilyl Chloride (Ph$_2$MeSiCl = 0.500 M) and Trimethylsilyl Chloride (TMS-Cl = 0.829 M) in Tetrahydrofuran at 25 °C

[NaPic], M × 10^3	k_2, L·mol^{-1} s^{-1}	
	Ph$_2$MeSiCl	TMS-Cl
17.90	$(7.39 \pm 0.19) \times 10^{-2}$	$(4.01 \pm 0.09) \times 10^{-1}$
8.95	$(7.76 \pm 0.24) \times 10^{-2}$	$(5.20 \pm 0.10) \times 10^{-1}$
4.47	$(7.92 \pm 0.36) \times 10^{-2}$	1.32 ± 0.05
2.24	$(8.54 \pm 0.33) \times 10^{-2}$	1.61 ± 0.03
1.12	$(9.87 \pm 0.33) \times 10^{-2}$	1.77 ± 0.11

Table VI. Second-Order Rate Constants for the Silylation of Potassium Picrate (KPic) with Triisopropylsilyl Chloride (i-Pr$_3$SiCl = 0.455 M) and *tert*-Butyldiphenylsilyl Chloride (t-BuPh$_2$SiCl = 0.375 M) in Acetonitrile with and without 18-Crown-6 as Addendum

[KPic], M × 10^3	[addendum], M × 10^3	k_2, L·mol^{-1} s^{-1}	
		i-Pr$_3$SiCl	t-BuPh$_2$SiCl
4.51		$(1.84 \pm 0.02) \times 10^{-3}$	$(1.38 \pm 0.09) \times 10^{-2}$
2.22		$(7.11 \pm 0.28) \times 10^{-3}$	$(9.74 \pm 0.40) \times 10^{-3}$
1.13		$(9.53 \pm 0.41) \times 10^{-3}$	$(8.46 \pm 0.68) \times 10^{-3}$
0.56		$(1.28 \pm 0.08) \times 10^{-2}$	$(1.03 \pm 0.08) \times 10^{-2}$
9.03	9.03	$(1.14 \pm 0.01) \times 10^{-2}$	$(7.08 \pm 0.31) \times 10^{-3}$
4.51	4.51	$(1.56 \pm 0.01) \times 10^{-2}$	$(1.15 \pm 0.03) \times 10^{-2}$
2.26	2.26	$(2.18 \pm 0.02) \times 10^{-2}$	$(1.83 \pm 0.04) \times 10^{-2}$
1.13	1.13	$(2.15 \pm 0.03) \times 10^{-2}$	$(2.12 \pm 0.03) \times 10^{-2}$
0.56	0.56	$(2.57 \pm 0.04) \times 10^{-2}$	$(1.80 \pm 0.06) \times 10^{-2}$

Table I provides a reference point for discussion of the silylation reaction in terms of the more familiar methylation of sodium phenolate in acetonitrile and tetrahydrofuran, with and without the addition of an amount of 15-crown-5 polybasic ether in an amount equivalent to that of the sodium phenolate. Reactions of unsubstituted sodium phenolate with silyl chlorides were too fast to measure. However, triphenylsilyl fluoride gave readily measurable rates by virtue of fluoride as a relatively poor leaving group. Table II compares rates for several substituted potassium phenolates with this reagent both in dimethyl sulfoxide, a good complexer for the potassium ion, and in acetonitrile, whose dielectric constant is similar to that of DMSO's but which is a much poorer cation complexer.

In order to compare the several silyl chlorides it was necessary to deactivate the phenolate ring with nitro groups. Tables III and IV compare the reactions of sodium 2,4-dinitrophenolate with *tert*-butyldiphenylsilyl chloride and triisopropylsilyl chloride in

Table VII. Second-Order Rate Constants for the Silylation of the Potassium Salt of Dibenzoylmethane (KDBM) with *tert*-Butyldiphenylsilyl Chloride (t-$BuPh_2SiCl$ = 0.221 M) and Triisopropylsilyl Chloride (i-Pr_3SiCl = 0.269 M) in Tetrahydrofuran at 25 °C

[KDBM], M × 10^3	k_2, $L\cdot mol^{-1}\,s^{-1}$	
	i-Pr_3SiCl	t-$BuPh_2SiCl$
3.77	$(3.37 \pm 0.27) \times 10^{-2}$	$(2.81 \pm 0.02) \times 10^{-2}$
1.89	$(7.44 \pm 0.28) \times 10^{-2}$	$(5.72 \pm 0.23) \times 10^{-2}$
0.94	$(6.85 \pm 0.13) \times 10^{-2}$	$(9.40 \pm 0.26) \times 10^{-2}$
0.47	$(8.45 \pm 0.24) \times 10^{-2}$	$(1.09 \pm 0.04) \times 10^{-1}$
0.24		$(1.05 \pm 0.14) \times 10^{-1}$

Table VIII. Conductance of Sodium Phenolate, Sodium 2,4-Dinitrophenolate, and Sodium Picrate in Tetrahydrofuran at 25 °C

	Λ, $\Omega^{-1}\cdot cm^2\cdot mol^{-1}$		
[phenolate], M	sodium phenolate	sodium 2,4-dinitrophenolate	sodium picrate
2.00×10^{-2}	6.30×10^{-3}		1.81×10^{-1}
1.00×10^{-2}	1.91×10^{-2}		1.88×10^{-1}
5.00×10^{-3}	4.22×10^{-2}	1.17×10^{-1}	2.20×10^{-1}
2.50×10^{-3}	8.66×10^{-2}	1.26×10^{-1}	2.82×10^{-1}
1.25×10^{-3}	1.46×10^{-1}	1.40×10^{-1}	3.71×10^{-1}
6.25×10^{-4}	2.27×10^{-1}	1.80×10^{-1}	5.15×10^{-1}

tetrahydrofuran and of sodium 2,6-di-*tert*-butyl-4-nitrophenolate with the same silylating agents in acetonitrile. Further nitration to the picrate ion leads to a sharp reduction in silylation rates, and Tables V and VI give the rate constants for the most reactive silyl chlorides, diphenylmethylsilyl chloride and trimethylsilyl chloride, with sodium picrate in tetrahydrofuran and potassium picrate with triisopropylsilyl chloride and *tert*-butyldiphenylsilyl chloride in acetonitrile with and without 18-crown-6 as addendum.

Finally, Table VII compares the potassium salt of dibenzoylmethane in its reactions with triisopropylsilyl chloride and *tert*-butyldiphenylsilyl chloride in tetrahydrofuran.

The trends in rate constants shown in Tables I–VII provide overwhelming evidence for the effect of ion aggregation on the reactivities of the alkali phenolates. However, conductance is a more fundamental tool than kinetics for quantitative evaluation of dissociation phenomena. Tables VIII and IX compare the conductance of sodium phenolate, 2,4-dinitrophenolate, and picrate in tetrahydrofuran and the effect of varied amounts of 15-crown-5 and Kryptofix 2·2·2 on the conductance of sodium 2,6-*di-tert*-butyl-4-nitrophenolate in tetrahydrofuran at 25 °C.

Although there was no means of comparing rates of reaction for all of the silylating agents with all of the phenolate salts, thermochemistry provided a straightforward means for comparison through heats of silylation. Table X gives a complete listing of heats of reaction for the four silyl chlorides with the substituted phenolate salts with sodium and potassium in acetonitrile and in THF. The enormous range of heats of reaction is consistent with the difficulty in finding a single set of conditions that would allow a complete kinetic comparison of all five silylating agents.

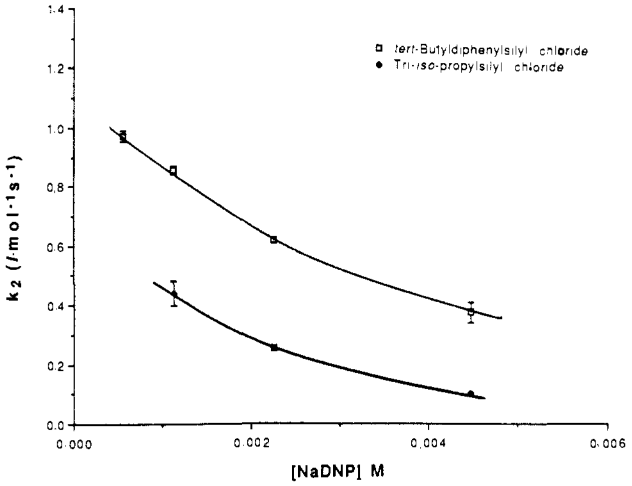

Figure 1. Plot of the calculated second-order rate constant for the silylation of sodium 2,4-dinitrophenolate with *tert*-butyldiphenylsilyl chloride and triisopropylsilyl chloride in tetrahydrofuran at 25 °C versus the molar concentration of the salt.

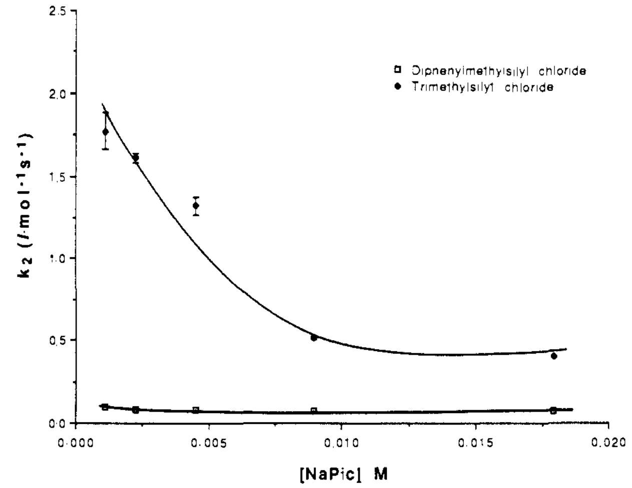

Figure 2. Plot of the calculated second-order rate constant for the silylation of sodium picrate with diphenylmethylsilyl chloride and trimethylsilyl chloride in tetrahydrofuran at 25 °C versus the molar concentration of the salt.

Discussion

A glance at Tables I–V suggest immediately that ion aggregation has an important effect on the rates of reaction of the alkali phenolates in tetrahydrofuran and acetonitrile. The reaction rates increase rapidly as the concentration decreases; addition of crown ethers or cryptands produce even greater rate increases. Rate constants determined under comparable conditions in acetonitrile (ϵ = 37.5) are greater than in tetrahydrofuran (ϵ = 7.6). Finally, ion aggregation is confirmed completely by the conductance measurements presented in Tables VIII and IX.

Table IX. Conductance of Sodium 2,6-Di-*tert*-butyl-4-nitrophenolate in Tetrahydrofuran at 25 °C with Varied Amounts of 15-Crown-5 and Kryptofix 2·2·2

			Λ, $\Omega^{-1}\cdot cm^2\cdot mol^{-1}$			
			crown			
[phenolate], M	Λ (no addenda), $\Omega^{-1}\cdot cm^2\cdot mol^{-1}$	[phenolate], M	0.5 equiv	1.0 equiv	2.0 equiv	Kryptofix, 2.0 equiv
3.00×10^{-2}	6.12×10^{-1}	3.00×10^{-2}	1.76	1.76	1.90	1.28×10^{1}
2.00×10^{-2}	5.13×10^{-1}	1.50×10^{-2}	1.56	1.66	1.81	1.30×10^{1}
1.50×10^{-2}	4.63×10^{-1}	7.50×10^{-3}	1.57	1.82	1.96	1.40×10^{1}
1.00×10^{-2}	4.37×10^{-1}	3.75×10^{-3}	1.80	2.16	2.38	1.60×10^{1}
5.00×10^{-3}	4.40×10^{-1}	1.88×10^{-3}	2.18	2.68	3.00	1.90×10^{1}
2.50×10^{-3}	5.07×10^{-1}	9.38×10^{-4}	2.73	3.38	3.86	2.32×10^{1}
1.25×10^{-3}	6.29×10^{-1}	4.69×10^{-4}	3.43	4.23	4.98	2.90×10^{1}
6.25×10^{-4}	8.93×10^{-1}	2.34×10^{-4}	4.17	5.24	6.38	3.59×10^{1}
3.13×10^{-4}	1.25					4.41×10^{1}
1.56×10^{-4}	1.73					

Table X. Heats of Reaction, $-\Delta H_{rxn}$, for Silylation of Substituted Sodium and Potassium Phenolates with Various Silylating Agents in Various Solvents at 25 °C[a]

	TMSiCl			$Ph_2MeSiCl$			t-$BuPh_2SiCl$	i-Pr_3SiCl
4-subst	A, K^+	A, Na^+	T, Na^+	A, K^+	A, Na^+	T, Na^+	A, Na^+	A, Na^+
H	37.14	28.24	21.32	41.36	30.31		24.62	22.23
	±0.40	±0.30	±0.30	±0.16	±0.45		±0.33	±0.06
F	35.23			39.70				
	±0.36			±0.41				
Cl	34.48			38.65				
	±0.39			±1.27				
CN	32.80			35.01				
	±1.11			±0.78				
NO_2	27.69	25.27		29.60	28.67		14.65	13.45
	±0.74			±1.40				
2,4-dinitro		15.57			17.44		11.33	8.46
		±0.19			±0.30		±0.15	±0.07
2,4,6-trinitro	11.27	9.77		12.90	11.54			
	±0.07	±0.77		±0.01	±0.11			
2,6-di-t-Bu-4-nitro	19.03	18.55	16.26	20.97	19.91	18.99		
	±0.25	±0.28	±0.09	±0.09	±0.09	±0.08		

[a] A, acetonitrile, T, THF.

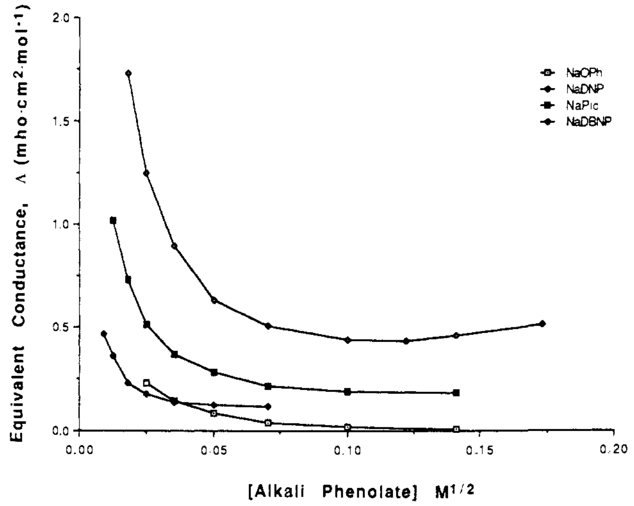

Figure 3. Plot of the equivalent conductances of sodium phenolate, sodium 2,4-dinitrophenolate, sodium picrate, and sodium 2,6-di-*tert*-butyl-4-nitrophenolate in tetrahydrofuran at 25 °C verus the square root of the molar concentration of the salt.

Figures 1 and 2 demonstrate graphically that it is futile to attempt any comparison of the reactivities of silylating agents unless there is a means to correct for varying ion aggregation as a function of concentration. Plots for the other rate data are not included here to conserve space, but they drive home the point just as emphatically. As we have shown in previous studies,[26a,31] the Acree equation

$$k_{obsd} = k_i\alpha + k_{agg}(1 - \alpha)$$

relates the observed second-order rate constant k_{obsd} to the rate constant for reaction of the free phenolate ions, k_i, and the separated rate for reaction of the aggregate or ion pair, k_{agg}. Although this treatment is applied customarily to the simple equilibrium between ions and ion pairs, it should be equally applicable to cases where the ion pair is replaced by a higher aggregate provided that there are no intermediate equilibria involving different states of aggregation with different reactivities as the concentration is changed. In every instance that we know of, the "free" ions are orders of magnitude more reactive than ion pairs and higher ion aggregates are probably, if anything, less reactive. As will be seen below, application of the Acree equation to the rate data tabulated here is for the most part surprisingly successful.

Conductance. Figure 3 is a conventional conductance plot of the data in Tables VIII and IX. If such a plot were linear, it would indicate that the cation and anion were completely dissociated across the range of concentration. However, curvature implies that an equilibrium exists between free ions and ion aggregates. Extrapolation of such a conductance plot to the y axis gives an approximate value for the limiting equivalent conductance Λ_0, the maximum conductance of the solvated ions at infinite dilution. However, in weakly associated systems such as that portrayed in Figure 3, accurate extrapolation of the conductance plot is very difficult.

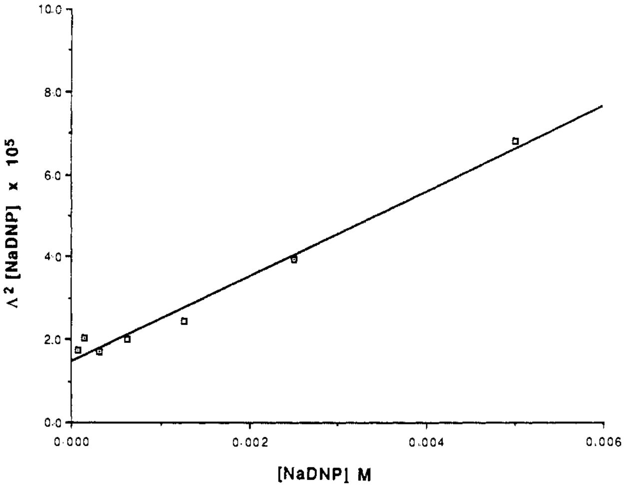

Figure 4. Wooster plot for sodium 2,4-dinitrophenolate in tetrahydrofuran at 25 °C.

Several mathematical treatments may be utilized to calculate the limiting equivalent conductance. The Marshall and Grunwald[32] and Shedlovsky[33] methods were applied but did not provide accurate values for Λ_0, which perhaps is not surprising since these methods require conductance data of the ionic species at very high dilution where interference from higher aggregates is negligible. As an alternative we applied the method of Wooster[34] based on triple ion formation as a model. Linearity of a Wooster plot indicates triple ion formation. The intercept of a Wooster plot is equal to $\Lambda_0^2K_d$, and if Λ_0 is known, then K_d can be calculated. Gilkerson[35a,b] has reported dissociation constants from Wooster plots for several alkali phenolates in tetrahydrofuran at 25 °C using estimates of 100 and 106 conductance units for the limiting equivalent conductance of sodium 2,4-dinitrophenolate and sodium picrate. K_d was calculated to be 9.3×10^{-10} and 1.05×10^{-8}

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[31] DePalma, V. M.; Arnett, E. M. *J. Am. Chem. Soc.* **1978**, *100*, 3514.

[32] Marshall, H. P.; Grunwald, E. *J. Chem. Phys.* **1953**, *21*, 2143.

[33] Shedlovsky, T. *J. Franklin Inst.* **1938**, *225*, 739.

[34] Wooster, C. B. *J. Am. Chem. Soc.* **1937**, *59*, 377.

[35] (a) Gilkerson, W. R.; Jackson, M. D. *J. Am. Chem. Soc.* **1979**, *101*, 4096. (b) Flora, H. B.; Gilkerson, W. R. *J. Phys. Chem.* **1973**, *77*, 1421.

Table XI. Parameters of Acree Plots for Silylation and Methylation of Alkali Phenolates in Tetrahydrofuran and Acetonitrile at 25 °C

phenolate	solvent	silyl halide and MeI	k_i, L·mol^{-1} s^{-1}	k_{agg}, L·mol^{-1} s^{-1}	correln coeff
Na 2,4-dinitro	THF	*t*-BuPh$_2$SiCl	5.52×10^2	1.31×10^{-1}	0.9533
Na 2,4-dinitro	THF	*i*-Pr$_3$SiCl	5.92×10^2	~0	0.9983
Na picrate	THF	Ph$_2$MeSiCl	9.58×10^2	6.50×10^{-2}	0.9818
Na picrate	THF	TMS-Cl	5.57×10^2	3.35×10^{-2}	0.9220
K picrate	MeCN	*i*-Pr$_3$SiCl	1.43×10^{-1}	~0	0.9752
Na 2,6-di-*t*-Bu-4-nitro	THF	*t*-BuPh$_2$SiCl	1.44	5.58×10^{-3}	0.9379
Na 2,6-di-*t*-Bu-4-nitro	THF	*i*-Pr$_3$SiCl	curved		
Na 2,6-di-*t*-Bu-4-nitro	THF	*i*-Pr$_3$SiCl + 15-crown-5 (1 equiv)	curved		
Na phenolate	THF	methyl iodide	2.3×10^7	~0	0.98

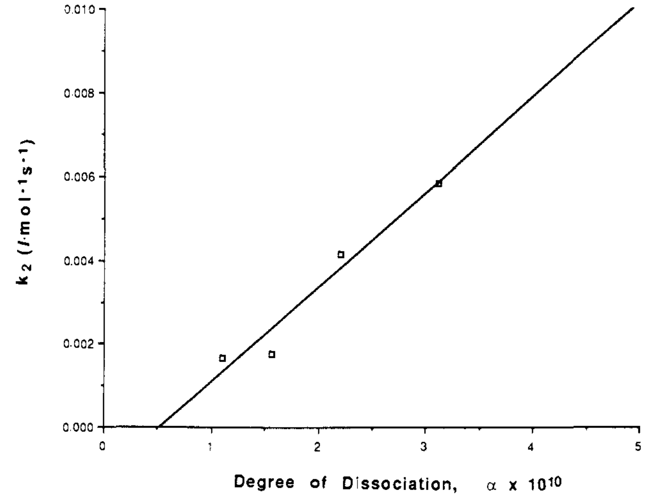

Figure 5. Acree plot for the alkylation of sodium phenolate with methyl iodide in tetrahydrofuran at 25 °C.

mol·L^{-1}, respectively, for these two salts. Application of these values of Λ_0 to the data in Table VIII gave dissociation constants of 1.49×10^{-9} and 1.27×10^{-8} for these two salts in reasonably close agreement with Gilkerson. Figure 4 is a Wooster plot for 2,4-dinitrophenolate in tetrahydrofuran, which is linear with a correlation coefficient of 0.9918. A similar treatment for sodium 2,6-di-*tert*-butyl-4-nitrophenolate yielded $K_d = 4.50 \times 10^{-8}$ mol·L^{-1}.

Acree Plots. Figure 5 shows an Acree plot for the methylation of sodium phenolate in THF which we have used as an external check on our procedures in view of the difficulties of studying silylation referred to above. Okamoto[36] has reported a dissociation constant of 2.17×10^{-22} mol for sodium phenolate in THF at 25 °C. This salt is very highly aggregated in this low polarity medium. α values were calculated from this K_d and applied to the data in Table I. Figure 5 shows good adherence to the Acree relationship with a correlation coefficient of 0.98, a slope of 2.27 $\times 10^7$ L·mol^{-1} s^{-1} and an intercept of −0.001. From this a rate constant for the free ion, $k_i = 2.3 \times 10^7$ L·mol^{-1} s^{-1}, may be calculated. In Table XI, all of the relevant data for Acree plots in this study are presented. Most of the correlation coefficients are reasonably high, although several of the plots led to uninterpretable curves perhaps because of changing degrees of aggregation over the concentration range studied.

Perusal of Table XI allows some conclusions about the relative reactivities of the free phenolate anions and the silylating agents. The high rate constant for reaction of unsubstituted phenolate anion with methyl iodide supports our inability to study the reaction by stopped-flow methods. In fact, Okamoto's dissociation constant indicates that the only reason we were able to obtain measurable rates with this system was because the very high degree of ion association had reduced the phenolate ion concentration to a miniscule level. Comparison of the calculated k_i values indicates that 2,4-dinitrophenolate is considerably more reactive than 2,6-di-*tert*-butyl-4-nitrophenolate against *tert*-butyldiphenylsilyl chloride. Steric hindrance is an obvious explanation.

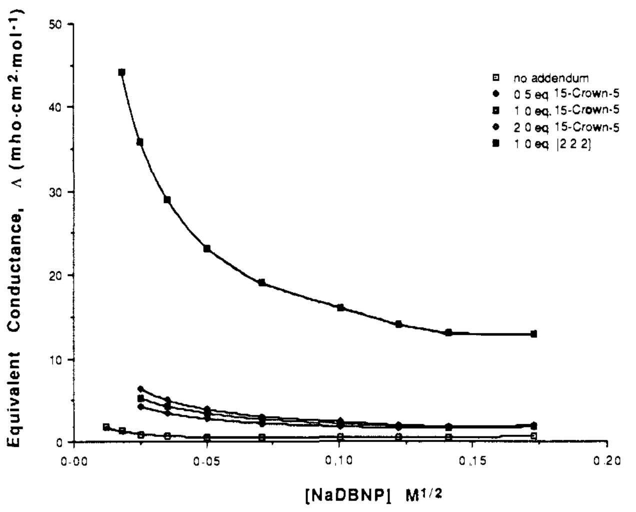

Figure 6. Plot of the equivalent conductance of sodium 2,6-di-*tert*-butyl-4-nitrophenolate in tetrahydrofuran at 25 °C with and without added crown ether and cryptand versus the square root of the molar concentration of the salt.

Table XII. Calculated Dissociation Constants for Sodium 2,6-Di-*tert*-butyl-4-nitrophenolate in the Presence of Crown Ether and Cryptand in Tetrahydrofuran at 25 °C

addendum	K_d,[a] mol·L^{-1}	rel K_d
none	4.50×10^{-8}	1
15-crown-5 (0.5 equiv)	1.72×10^{-7}	3.8
15-crown-5 (1.0 equiv)	6.22×10^{-7}	13.8
15-crown-5 (2.0 equiv)	8.73×10^{-7}	19.4
Kryptofix 2·2·2 (1.0 equiv)	3.16×10^{-5}	702

[a] Calculated from Λ_0 = 100mho·cm^2·mol^{-1}.

Direct comparison of the four silyl chlorides through their rates of reaction with picrate and 2,4-dinitrophenolate ions cannot be done with the data in Table XI. However, the fact that diphenylmethylsilyl chloride and trimethylsilyl chloride give rates with the highly deactivated picrate ion that are comparable to those of *tert*-butyldiphenylsilyl chloride and triisopropylsilyl chloride with 2,4-dinitrophenolate suggests that the first two silylating agents are much more reactive than the latter two. Finally, as might be expected, in all cases the k_i values are much larger than those for k_{agg}.

Influence of Additives. Tables I, III, and IX indicate the effect that 15-crown-5 and Kryptofix-2·2·2 have on favoring the conductance and rates of reaction for phenolate anion. Figure 6 portrays graphically the effect of adding increasing amounts of crown ether on the conductance of sodium 2,6-di-*tert*-butyl-4-nitrophenolate in THF. Table XII shows how the additives affect the calculated dissociation constant, especially the dramatic difference between the crown and cryptand types of ligand. These are based on an estimated Λ_0 of 100 for the 2,6-di-*tert*-butyl-4-nitrophenolate anion from Wooster plots[34,35] of the conductance data. The dramatic difference between the effectiveness of the cryptand and the crown ether is clear. In tetrahydrofuran, phenolates are probably associated with the complexed sodium cation as shown in Figure 7. As shown here, the complexed sodium ion is still capable of forming triple ions with the anion.

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[36] Okamoto, K.; Kinoshita, T.; Oishi, T.; Moriyama, T. *J. Chem. Soc., Perkin Trans. 2* **1978**, 453.

Table XIII. pK_a's, Hammett Substituent Parameters, Chemical Shifts of Hydroxyl Protons of Substituted Phenols in Dimethyl Sulfoxide, Heats of Reaction of Triphenylsilyl Fluoride with Potassium Phenolates in Dimethyl Sulfoxide and Acetonitrile, and Corresponding Second-Order Reaction Rate Constants All at 25 °C

4-substituted phenols	pK_a		σ, σ^-	δ(DMSO)	ΔH, kcal mol^{-1}		k_2, L·mol^{-1} s^{-1}	
	MeCN	DMSO			DMSO	MeCN	DMSO	MeCN
methoxy		19.08	−0.16	8.90	23.03 ±0.78			
t-Bu	27.48	19.05	−0.17	9.10	21.84 ±0.66		1.28 ± 0.06 × 10^3	
H	26.6	17.96	0	9.30	20.68 ±0.17	22.97 ±1.39		
fluoro		18.14	0.17	9.35	22.40 ±0.84	21.88 ±0.43		
phenyl		17.14	0.11	9.54	18.07 ±0.55	21.06 ±0.74	2.10 ± 0.34 × 10^2	
chloro	25.44	16.75	0.27	9.65	18.80 ±0.09		6.61 ± 1.04 × 10^2	1.60 ± 0.43
trifluoromethyl		15.26	0.65	10.30		17.50 ±0.61	3.23 ± 0.14 × 10^2	2.91 ± 0.02
carbomethoxy		14.33	0.64	10.33	10.71 ±0.13		1.72 ± 0.05 × 10^2	
cyano	22.7	13.18	0.88	10.60	9.58 ±0.32		7.58 ± 0.38	1.16 ± 0.18
nitro	20.7	10.78	1.24	11.05	5.58 ±0.14		2.21 ± 0.07 × 10^{-1}	7.16 ± 0.58 × 10^{-1}

Table XIV. Correlations of Various Properties of Substituted Phenols and Their Salts in Dimethyl Sulfoxide from Table XIII

	Y axis				
x axis	pK_a	σ, σ^-	δ	ΔH(DMSO)	log k_2(DMSO)
pK_a	X				
σ, σ^-	(10 pts) corr = −0.9863 m = −5.7299 b = 18.2469	X			
δ	(10 pts) corr = −0.9866 m = −3.7851 b = 53.3061	(10 pts) corr = 0.9905 m = 0.6541 b = −6.0546	X		
ΔH_{rxn}	(9 pts) corr = 0.9865 m = 0.4417 b = 8.8728	(9 pts) corr = −0.9644 m = −0.0731 b = 1.5549	(9 pts) corr = −0.9864 m = −0.1125 b = 11.6415	X	
log k_2	(7 pts) corr = 0.9142 m = 1.8941 b = 11.6356	(7 pts) corr = −0.8695 m = −0.3149 b = 1.1118	(7 pts) corr = −0.8524 m = −0.4355 b = 10.9039	(6 pts) corr = 0.8980 m = 4.0153 b = 6.9293	X

In Table VI, the addition of 18-crown-6 to the silylation reactions of potassium picrate in acetonitrile gives a minimal affect on its reaction with triisopropylsilyl chloride and *tert*-butyldiphenylsilyl chloride. This system should represent the maximum opportunity for dissociation in view of the size of the potassium ion, delocalization in the picrate ion, and the relatively high dielectric constant of the medium.

Heats of Reaction. The above discussion documents in detail the difficulties of making direct comparisons between the reactivities of the silyl chlorides and the various phenolate ions. Fortunately, reaction calorimetry provides a straightforward approach to the determination of meaningful thermochemical comparisons. Table X provides heat of reaction data for all of the silyl chlorides with sodium phenolate in acetonitrile and also heats of reaction for a number of silyl chlorides reacting with substituted sodium and potassium phenolates in acetonitrile and tetrahydrofuran. These allow comparisons of the effects of counterion, solvent, and steric effects on the silylation reaction.

Electronic influences are much more clearly displayed in Table XIII, where heats of reaction in DMSO and acetonitrile are comparable directly with the second-order rate constants for the series of para-substituted potassium phenolates and also with pK_a data in both solvents, the chemical shifts of the hydroxyl protons of the parent phenols, and the Hammett σ and σ^- parameters for the phenolate anions, all at 25 °C. Heats of reaction were obtained by titration of the silylating agent (0.1–0.2 mol) in the solvent stated into a calorimeter containing the alkali phenolate solution in a concentration of 5–10 mmol. In all cases shown in Table X, heats of reaction with a potassium phenolate were larger than for the sodium salt. The difference was most pronounced for the phenolate ions that were most dissociated by virtue of substitution with nitro groups and steric hindrance by *tert*-butyl groups at the

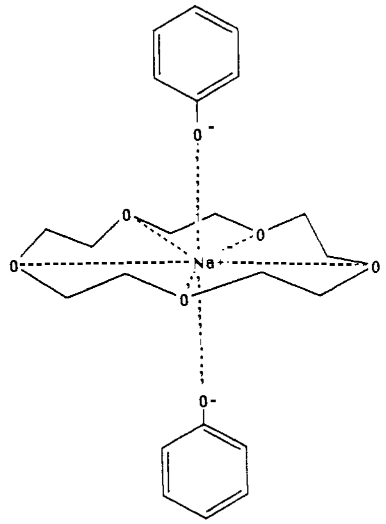

Figure 7. Proposed structure of a triple ion.

2,6-positions. Cation effects, which presumably reflect differences in ion aggregation, were enormous (10–11 kcal/mol) for unsubstituted potassium and sodium phenolate. However, such a large effect is unprecedented and should be regarded with suspicion, although it was repeated many times, always with the same result. Again, the few data that allow comparison between acetonitrile and the low dielectric solvent, THF, presumably reflect the large endothermic price for separating cation and anion that must be paid in opposition to the large exothermic heats of reaction of the silyl chlorides with the phenolate anions.

Looking across Table X, heats of reaction with sodium phenolate provide the only complete comparison of the reactivities of the silyl chlorides which clearly fall in the order diphenylmethyl > trimethyl > *tert*-butyldiphenyl > triisopropyl. The same order is found for reaction with sodium *p*-nitrophenolate, and although the rest of the data are incomplete, there are no exceptions to the above order. Clearly, steric effects dominate the reactivities of the four silyl chlorides, although the superior reactivity of $Ph_2MeSiCl$ over that of Me_3SiCl probably reflects the greater electron-withdrawing power of phenyl over methyl. Steric hindrance at the 2- and 6-positions also reduces the reactivity of the phenolate anion as shown by the drastically reduced heat of reaction of 2,6-di-*tert*-butyl-4-nitro- when compared to unsubstituted 4-nitrophenolate anion. This effect is so large that we were unable to obtain good heats of reaction for the 2,6-disubstituted compounds with *tert*-butyldiphenylsilyl chloride or triisopropylsilyl chloride, the most sterically hindered silylating agents. Steric hindrance also contributed to differences in rate versus concentration profiles of reactions.

Steric effects reported here support the observations of Lasocki and Kowalski[19] on the reactions of alkyl-substituted silylamides with *p*-nitrophenol in dioxane. They found the order of reactivity to be $Me_3 > Me_2Et > Et_3$. Although the reaction rates for the silyl chlorides cover too wide a range for a complete comparison, the results presented in Tables III–V and Figures 1 and 2 support the steric effects on enthalpies of reaction described above.

Electronic Effects. Table XIII provides a nearly complete comparison of electronic effects on the chemical shift of the phenol proton, the heats, and second-order rate constants for reaction of the 4-substituted potassium phenolates in DMSO and acetonitrile with triphenylsilyl fluoride and compares these with their pK_a's in both of these solvents and their Hammett parameters. These comparisons of rates and heats of reaction were made possible by exploiting the relatively poor leaving group ability of fluoride compared to the much more reactive chlorides. Even with that restriction, reaction rates were too fast to measure by stopped-flow for potassium *p*-methoxyphenolate.

With only a few minor exceptions, excellent proportionality among all the different properties is obvious from perusal of Table XIII. Rather than present an extensive series of linear correlation plots, we will simply tabulate the results of a number of extrathermodynamic correlations that we have made in Table XIV. In view of the good linearity for most of these correlations, it is reasonable to assume that other types of correlations that we have not made among these properties would also give linear correlations.

Conclusions. A kinetic and thermodynamic investigation of the silylation reactions of alkali phenolates with several silyl halides in tetrahydrofuran, acetonitrile, and dimethyl sulfoxide was undertaken to establish quantitative comparisons of the reactivities of the silyl halides. Reaction rates and heats were strongly dependent on solvent polarity, ion pairing of the alkali phenolates, and steric and the electronic influences of substituents. The dissociation constants were obtained from Wooster plots of the conductance data for several of the salts, thus demonstrating the presence of triple ions or higher aggregates. The dissociation constants were used to correct the concentration effects on reaction rates through the Acree equation which allowed comparison of the reactivities of the silyl halides with the phenolate ions when freed from their counterions.

Reaction rates of the silyl chlorides covered such a range that it was not possible to make direct comparisons between them by our kinetic techniques; however, heats of reaction with sodium phenolate in acetonitrile demonstrated clearly the overriding effect of steric hindrance in the following order: diphenylmethyl > trimethyl > *tert*-butyldiphenyl > triisopropyl.

Less reactive triphenylsilyl fluoride gave measurable second-order rate constants and also heats of reaction with most of the alkali phenolates. Comparison of these results with the chemical shifts of the phenolic proton, with the pK_a's of the phenols, and with the Hammett substituent constants shows that reactivities of the 4-substituted phenolate ions with triphenylsilyl fluoride are controlled by electronic factors. Rates of reaction of sodium phenolate with methyl iodide in tetrahydrofuran and acetonitrile are several orders of magnitude slower than are comparable rates for reactions with the silyl chlorides.

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