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Equilibrium Lithium Ion Pair Acidities of 2-Phenyl-1,3-dithiane and 2-*p*-Biphenylyl-1,3-dithiane: A New Approach[1]

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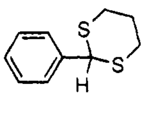
A new technique has been developed to measure the equilibrium ion pair acidities of 2-phenyl-1,3-dithiane (**1**) and 2-*p*-biphenylyl-1,3-dithiane (**2**) with lithium as counterion. The lithium ion pair p*K* values ($pK_{Li/THF}$) in tetrahydrofuran (THF) at 25 °C were determined to be 29.4 (**1**, contact ion pair (CIP)), 28.2 (**2**, CIP), and 29.3 (**2**, solvent separated ion pair (SSIP)), as compared to the values of 30.5 (**1**, CIP) and 29.1 (**2**, CIP) measured previously with cesium as counterion. At 25 °C, the lithium salt of 2-*p*-biphenylyl-1,3-dithiane (**2L**) exists as a mixture of CIP, SSIP, and free ion, with the ratio of [CIP]/[SSIP] = 11.4. The results are consistent with the stronger electrostatic interactions between lithium cation and the dithianyl anion for CIP than for SSIP, resulting in a lower pK_a for CIP. The equilibrium constant of SSIP $\rightleftharpoons$ CIP for **2L** was measured at 25 to −20 °C to give $\Delta H^\circ = 1.43$ kcal mol^{-1}, $\Delta S^\circ = 9.6$ eu. The SSIP of **2L** was found to dissociate significantly to the free ion in dilute THF solution. The measured dissociation constants of **2L** to the free ion at 25 °C are $(1.22 \pm 0.08) \times 10^{-6}$ M and $(1.39 \pm 0.09) \times 10^{-5}$ M for CIP and SSIP, respectively. The slow transmetalation reactions of **1** and **2** with different lithium indicator anions also allowed the measurements of some kinetic rate constants at 25 °C, from which Brønsted β values were obtained for **1** and **2**.

Since their introduction in 1965 by Corey and Seebach,[2] 1,3-dithianes and their lithium salts have found extensive applications in organic synthesis. These include the use of 1,3-dithianyllithium in the synthesis of many natural products such as sesquiterpenoids,[3] insect pheromones,[4] and a number of other natural products in recent years.[5] One of the most remarkable examples is the synthesis of several members of the steganin family by the exclusive use of 1,3-dithianyllithium.[6] A number of review articles on the synthetic applications of 1,3-dithianes have been published.[7]

Since the reactions involving 1,3-dithianyl salts are usually carried out in tetrahydrofuran (THF) solvent, determination of properties such as ion pair acidities of 1,3-dithianes (which, in turn, reflect the thermodynamic stability of their conjugate anions) and types of ion pairing of dithianyl salts in THF should aid understanding. Studies in the past have pointed out that reaction rates or chemical outcomes might be significantly affected by different types of ion pairs or free ions in solution.[8] We have previously reported the pK_a values for a number of 1,3-dithianes with cesium as counterion in THF,[9] where all dithianylcesium salts were shown to exist as contact ion pairs. pK_a measurements of some 1,3-dithianes have also been reported for cesium salts in cyclohexylamine (CHA)[10,11] from this laboratory, for lithium salts in THF by Fraser,[12] and for free ions in dimethyl sulfoxide (DMSO) by Bordwell.[13] Because of the relatively high solution concentrations used by Fraser[12] in the ^{13}C-NMR measurements, the pK_a values in this case may be affected by the aggregation of the lithium salts in THF. This is particularly true for those dithiane anions whose charge is highly localized. We thus set out to measure the pK_a values of 2-phenyl-1,3-dithiane (**1**) and 2-*p*-biphenylyl-1,3-dithiane (**2**) in THF, using the UV−vis spectroscopic technique described previously.[9] We expect the lithium salts of both **1** and **2** to be monomeric in THF, since the lithium salt of **1** was found to be monomeric in the solid state by Amstutz, Dunitz, and Seebach.[14]



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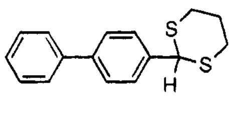

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Table 1. Spectral Data for Lithium and Cesium Salts of 1 and 2 in THF at 25 °C

compds	counterion	λ_{max} (nm)	ϵ (cm^{-1} M^{-1})
1	Cs	365	14 590
1	Li	322	13 100
2	Cs	456	35 439
2	Li	388 (CIP)	27 500[a]
		485 (SSIP)	27 500[a]
		506 (free ion)	27 500[a]

[a] The calculation was based on the assumption that the extinction coefficients are identical for CIP, SSIP, and free ion.

Results and Discussion

Ion Pairing. Table 1 lists the maximum absorption wavelength (λ_{max}) of the UV–vis spectra for the lithium and cesium salts of **1** and **2**, together with their corresponding molar absorptivity. Comparisons of the difference in λ_{max} between the corresponding lithium and cesium salts indicate that **1L** (the lithium salt of **1**) exists as contact ion pair (CIP) in THF solution, while **2L** (the lithium salt of **2**) is a mixture of contact ion pair (CIP, λ_{max} = 388 nm), solvent-separated ion pair (SSIP, λ_{max} = 485 nm), and free ion (λ_{max} = 506 nm). It is interesting to note the enormous difference of nearly 100 nm between the λ_{max} of CIP (388 nm) and that of SSIP (485 nm) for **2L**! Even though the negative charge is somewhat stabilized by the phenyl and biphenylyl groups in **1L** and **2L**, respectively, the delocalization of the charge is apparently not extensive enough to preclude the dominant formation of contact ion pairs. The extra anion stabilization from the *p*-phenyl group in **2L**, on the other hand, does promote the formation of SSIP and its dissociation to the free ion. Abatjoglou, Eliel, et al.[15] have previously reported in their NMR study that 2-lithio-2-phenyl-1,3-dithiane (**1L**) exists primarily as CIP in THF, while the addition of hexamethylphosphoramide (HMPA) converts CIP to SSIP.

Dissociation of 2L to Free Ions. The UV–vis spectra of **2L** at different concentrations are shown in Figure 1. The λ_{max} at 388 nm (peak A) due to CIP is not affected by the changing concentrations, but the λ_{max} of peak B shifts to the longer wavelength as the solution is diluted. Furthermore, the relative ratio of the absorbance at peak A (λ_{max} = 388 nm) to the absorbance at the λ_{max} of peak B decreases as the **2L** concentration is reduced. These observations are consistent with a free ion contribution to peak B at dilute concentrations. Indeed, peak B was later shown to be the mixed contribution from SSIP (λ_{max} = 485 nm) and the free ion (λ_{max} = 506 nm). Thus, the dissociation of CIP and SSIP to the free ion becomes more significant at lower concentrations, resulting in a reduced [CIP]/{[SSIP] + [free ion]} ratio.

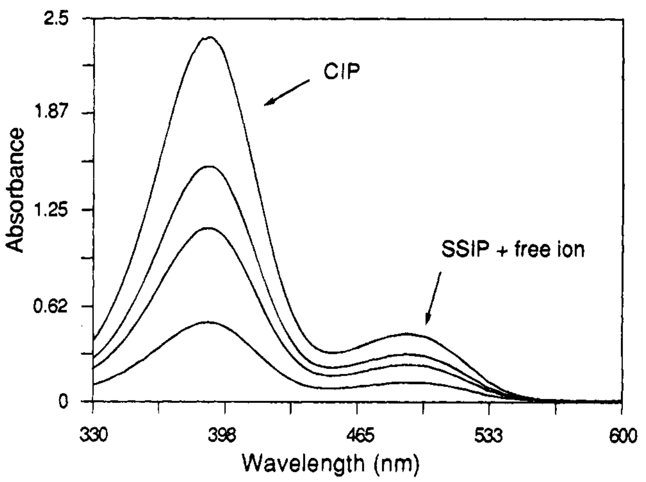


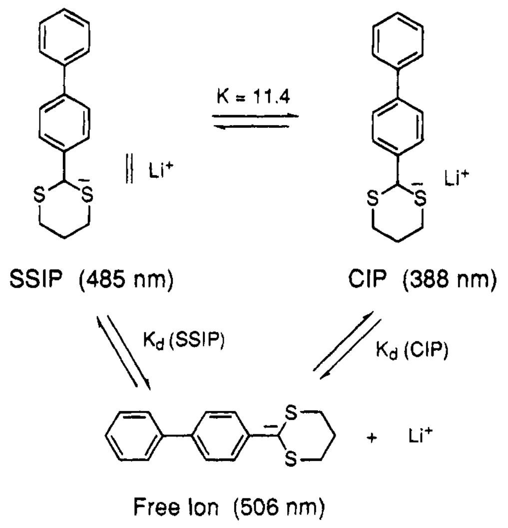
Figure 1. Concentration dependence of spectra of **2L** in THF at 25 °C.

Table 2. Dissociation Constants of 2L to Free Ion at 25 °C in THF

10^{-4} M [CIP][a]	10^{-5} M [free ion][b]	10^{-6} M $K_{d(CIP)}$ = [free ion]^2/[CIP]
1.97	1.60	1.30
2.07	1.53	1.13
2.65	1.71	1.10
3.57	2.17	1.32
4.82	2.42	1.22
5.22	2.53	1.23
6.94	2.89	1.20
7.29	3.03	1.26
average		(1.22 ± 0.08) × 10^{-6} M

[a] The corresponding SSIP concentrations can be calculated by [SSIP] = [CIP]/11.4. [b] The thermodynamic cycle for the three species is outlined in Scheme 1.

Scheme 1. Thermodynamic Cycle for the Equilibrium of CIP, SSIP, and the Free Ion for 2L



The presence of the free ion was further confirmed by common ion effects. The addition of an excess amount of lithium tetraphenylborate to varying concentrations of **2L** shifts peak B to 485 nm which is attributed to SSIP. A constant ratio of [CIP]/[SSIP] = 11.4 at 25 °C was obtained from these spectra where the free ion contribution is negligible. Subtraction of the spectrum of CIP and SSIP from the mixed spectrum of CIP, SSIP, and the free ion gives a peak with the absorption maximum at 506 nm due to the free ion. The relative amount of CIP and the free ion at different **2L** concentrations is listed in Table 2, along with the calculated dissociation constants of CIP to the free ion, $K_{d(CIP)}$, as defined in Scheme 1. The average $K_{d(CIP)}$ for the dissociation of CIP to the free ion is calculated to be (1.22 ± 0.08) × 10^{-6} M from Table

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Table 3. Dependence of λ_{max} for the Salts of 2 in THF on the Cation Radius

cation	r (Å)	λ_{max} (nm)
Li^+	0.60	388
Na^+	0.96	422[a]
Cs^+	1.66	455
M^+ (SSIP)	~4.5	485
free ion		506

[a] The spectrum for the sodium salt of **2** was obtained by adding an excess amount of sodium perchlorate to a solution of **2C**.

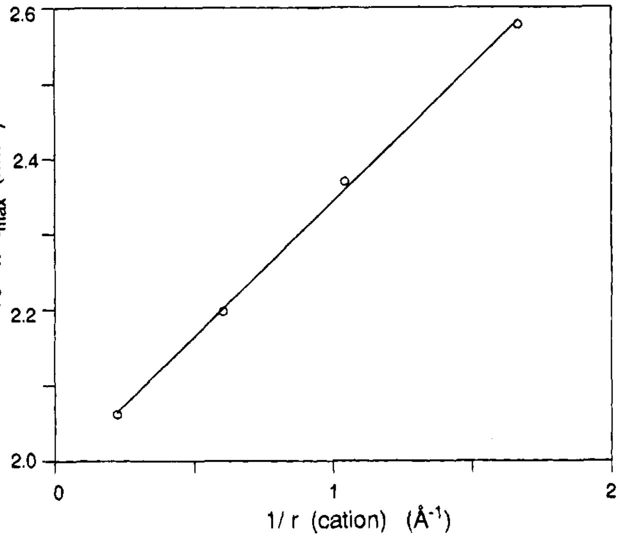

Figure 2. Correlation with $1/r$ of the wavenumber for the ion pairs and free ion of **2**. The data are taken from Table 3.

2, and the corresponding dissociation constant for SSIP to the free ion is $K_{d(SSIP)} = K_{d(CIP)} \times 11.4 = (1.39 \pm 0.09) \times 10^{-5}$ M. The magnitude of these dissociation constants is comparable to that of fluorenyllithium and related organolithium compounds.[16] Dissociation is less significant for CIP due to the stronger electrostatic interactions between lithium cation and the dithianyl anion. On the other hand, the dissociation of SSIP to the free ion is substantial at dilute concentrations. For an assumed total **2L** concentration of 2×10^{-4} M in THF, the relative amounts of CIP, SSIP, and free ion at 25 °C are 1.70×10^{-4}, 1.49×10^{-5}, and 1.44×10^{-5} M (or 85%, 7.5%, and 7.5%), respectively (almost equal amounts of SSIP and of the free ion!).

It has been well documented that the absorption bands of the fluorenyl ion pairs shift to the longer wavelength as the radius of the cation increases. In fact, the inverse of cation radius is linearly related to the wavenumber of the transition.[17] Table 3 lists the radius for several cations and our experimentally observed maximum absorption wavelength (λ_{max}) for the ion pairs and the free ion of **2L**. The plot of $1/r_{cation}$ vs $1/\lambda_{max}$ shows a linear relationship, as illustrated in Figure 2. Extrapolation to $1/r_{cation} = 0$ yields $\lambda_{max} = 507.6$ nm for the free ion of **2L**, close to the experimental value of 506 nm.

$\Delta H°$ and $\Delta S°$ for the Equilibrium SSIP ⇌ CIP of 2L. In the presence of an excess amount of lithium tetraphenylborate, the equilibrium constant for SSIP ⇌ CIP can be conveniently measured without interference from the free ion. The equilibrium constant was measured at several temperatures from 25 to −20 °C, and the results are summarized in Table 4. Figure 3 shows the Van't Hoff plot for ln K vs $1/T$, from which $\Delta H°$ and $\Delta S°$ were calculated to be 1.43 kcal mol^{-1} and 9.6 eu, respectively. SSIP is clearly favored at low temperature, and the positive entropy change is consistent with more free solvent molecules in going from SSIP to CIP.

Table 4. Temperature Dependence of the Equilibrium Constant for SSIP ⇌ CIP of 2L in THF

temp (°C)	K = CIP/SSIP	ln K
24.8	11.4	2.434
11.7	10.0	2.305
−3.8	8.67	2.160
−13.5	7.86	2.062
−20.4	7.33	1.992

Ion Pair Acidities. Ion pair acidity refers to the transmetalation equilibrium in eq 1 where RH is the

$$RH + In^-M^+ \overset{K}{\rightleftharpoons} InH + R^-M^+ \tag{1}$$

$$-\log K = pK_{RH} - pK_{InH}$$

carbon acid whose pK_a is to be measured and InH is an appropriate indicator whose pK is on the established acidity scale.[16a,18] The acidity scale was built up by measuring the successive transmetalation equilibrium constants spectroscopically with the UV−vis technique described previously. For cesium as countercation (i.e., $M^+ = Cs^+$), the established acidity scale covers compounds whose pK's range from about 12 to 39 with fluorene as the standard assigned its pK in DMSO, 22.9 (per hydrogen).[19] Our acidity scale with lithium as counterion, however, is limited to compounds with pK's lower than 24.[20] For compounds with pK's in the mid to high 20's, such as our title compounds **1** and **2**, the proton−lithium cation exchange in eq 1 becomes extremely slow and renders the direct measurement of the equilibrium constants almost impossible without substantial decomposition of the lithium salts. This limitation has prompted us to seek other methods of measuring the acidity of compounds such as **1** and **2**.

Although experimentally impractical, eqs 2 and 3 (where Fl stands for fluorenyl) define the difference between the pK's of our compound RH and the anchor point fluorene with lithium and cesium as counterions, respectively. Since both pK's of the fluorenyllithium and

$$RH + Fl^-Li^+ \overset{K_1}{\rightleftharpoons} FlH + R^-Li^+ \tag{2}$$

$$-\log K_1 = pK_{RH(Li)} - pK_{FlH}$$

$$RH + Fl^-Cs^+ \overset{K_2}{\rightleftharpoons} FlH + R^-Cs^+ \tag{3}$$

$$-\log K_2 = pK_{RH(Cs)} - pK_{FlH}$$

fluorenylcesium are taken as 22.9, eq 4 follows. We have

$$-\log(K_1/K_2) = pK_{RH(Li)} - pK_{RH(Cs)} \tag{4}$$

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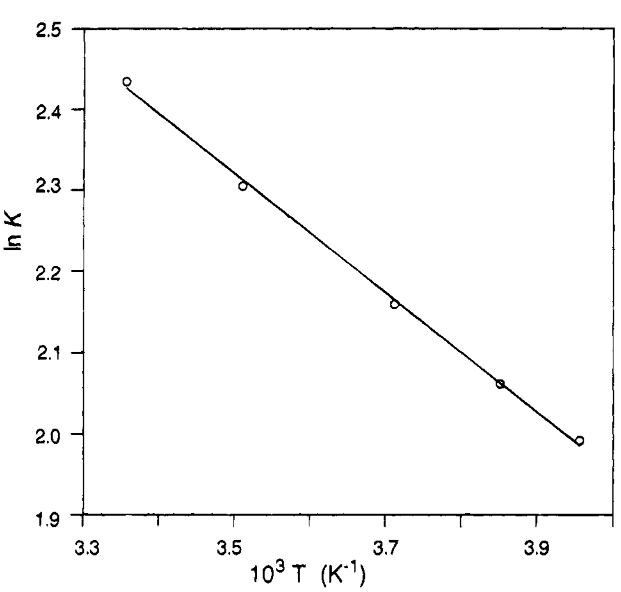

Figure 3. Van't Hoff plot for the equilibrium of SSIP ⇌ CIP for **2L**.

previously reported the cesium ion pair acidity ($pK_{RH(Cs)}$) for both compounds **1** and **2**,[8] so the determination of the ratio K_2/K_1 would lead to the desired $pK_{RH(Li)}$. Subtraction of (3) from (2) yields (5), where dissociation of all ion pair species to their free ions constitutes a thermodynamic cycle.

$$\begin{tikzcd}
R^-Cs^+ + Fl^-Li^+ \arrow[r, leftrightarrow, "K_3"] \arrow[d, leftrightarrow, "K_{d(FlLi)}K_{d(RCs)}"'] & Fl^-Cs^+ + R^-Li^+ \arrow[d, leftrightarrow, "K_{d(FlLi)}K_{d(RLi)}"] \\
R^- + Cs^+ + Fl^- + Li^+ \arrow[r, leftrightarrow, "K_4 = 1"] & Fl^- + Cs^+ + R^- + Li^+
\end{tikzcd} \tag{5}$$

From this thermodynamic cycle, eq 6 follows

$$K_1/K_2 = K_3 = \frac{K_{d(FlLi)}K_{d(RCs)}}{K_{d(FlCs)}K_{d(RLi)}} \tag{6}$$

where K_d's represent the dissociation constants of Fl^-Li^+, Fl^-Cs^+, R^-Li^+, and R^-Cs^+ to the corresponding free ions. The direct measurement of equilibrium constant K_3 is rather difficult because of the complicated spectral overlap of all four ion pair species. On the other hand, the measurements of all four dissociation constants could ultimately lead to the ion pair acidities of **1** and **2**. The dissociation constants for Fl^-Li^+ and Fl^-Cs^+ are known from conductivity studies.[16] Similar measurements could presumably be made for R^-Li^+ and R^-Cs^+, but we have developed a new method that avoids these lengthy conductivity studies.

In the course of this study, we discovered that cationic exchange between **1L** or **2L** with cesium tetraphenylborate is surprisingly fast, as described by eq 7. Thus, we

$$\begin{tikzcd}
Cs^+BPh_4^- + R^-Li^+ \arrow[r, leftrightarrow, "K_5"] \arrow[d, leftrightarrow, "K_{d(CsBPh_4)}K_{d(RLi)}"'] & R^-Cs^+ + Li^+BPh_4^- \arrow[d, leftrightarrow, "K_{d(LiBPh_4)}K_{d(RCs)}"] \\
BPh_4^- + Cs^+ + R^- + Li^+ \arrow[r, leftrightarrow, "K_6 = 1"] & Fl^- + Cs^+ + BPh_4^- + Li^+
\end{tikzcd} \tag{7}$$

observed the rapid change in the spectrum of **1L** or **2L** to a mixed spectrum due to (**1L** + **1C**) or (**2L** + **2C**) upon addition of cesium tetraphenylborate to a solution of **1L** or **2L**. A similar change was also observed when lithium tetraphenylborate was added to a solution of **1C** or **2C**. The substantial difference in the λ_{max}'s (Table 1) of the lithium and cesium salts allows ready deconvolution of the mixed spectra (Figure 4; see acidity measurements in the Experimental Section for more details) and, hence, the measurement of the equilibrium constant K_5. K_5 is related to the dissociation constants, K_d's, for $Li^+BPh_4^-$, $Cs^+BPh_4^-$, R^-Li^+, and R^-Cs^+ by eq 8:

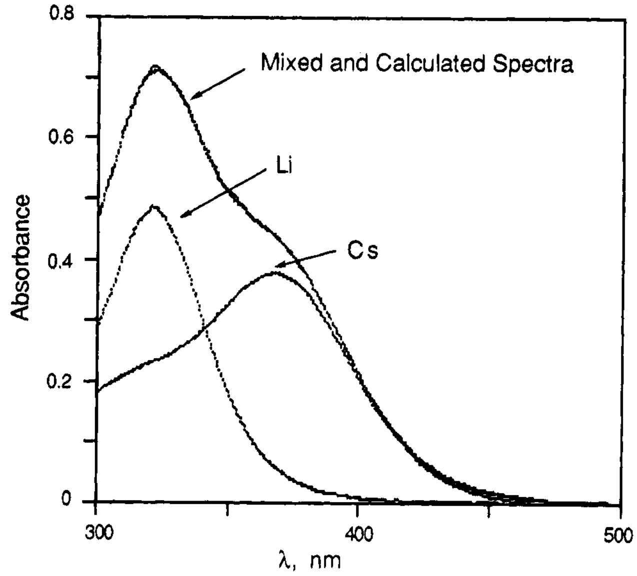

Figure 4. An example of the linear least-squares deconvolution of the mixed spectra of **1L** and **1C** at 25 °C. Note the excellent overlap of the mixed and the calculated spectra.

Table 5. Dissociation Constants[a] for Fluorenyllithium, Fluorenylcesium, Lithium Tetraphenylborate, and Cesium Tetraphenylborate in THF at 25 °C

salts	K_d (M)
Fl^-Li^+	3.89×10^{-6}
Fl^-Cs^+	1.38×10^{-8}
$Li^+BPh_4^-$	5.1×10^{-5}
$Cs^+BPh_4^-$	1.49×10^{-6}

[a] The dissociation constants were taken from ref 16.

$$K_5 = \frac{K_{d(CsBPh_4)}K_{d(RLi)}}{K_{d(LiBPh_4)}K_{d(RCs)}} \tag{8}$$

Combining eqs 4, 6, and 8 yields eq 9

$$pK_{RH(Li)} = pK_{RH(Cs)} - \log\left(\frac{K_{d(FlLi)}K_{d(CsBPh_4)}}{K_{d(FlCs)}K_{d(LiBPh_4)}}\right) + \log K_5 \tag{9}$$

where $pK_{RH(Cs)}$ was reported by us previously,[9] K_5 can be measured by the spectroscopic method already described, and all four dissociation constants, listed in Table 5, have been reported by other authors employing conductivity studies.[8c,16] The acidity of **1** or **2** can thus be calculated from eq 9; the results are tabulated in Table 6.

The pK_a values for **2** with CIP and SSIP as the conjugate base are 28.2 and 29.3, respectively. The difference in these values reflects the stronger electrostatic interaction between lithium cation and the dithianyl

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Table 6. pK_a's for Compounds 1 and 2 in THF with Lithium as Counterion at 25 °C

	1	2
p$K_{RH(Cs)}$	30.5	29.1
K_5	1.7[a]	1.0[a,b]
p$K_{RH(Li)}$[c]	29.4 (CIP)	28.2 (CIP) 29.3 (SSIP)[d]

[a] The uncertainty associated with K_5 is estimated to be about 15%. [b] This equilibrium constant is for the equilibrium between **2L** contact ion pair and cesium tetraphenylborate, as depicted in eq 7. [c] The uncertainty associated with p$K_{RH(Li)}$ is estimated to be ±0.2 pK units. [d] p$K_{(SSIP)}$ = p$K_{(CIP)}$ − log(11.4) where 11.4 is the equilibrium constant for SSIP $\rightleftharpoons$ CIP for **2L**, as shown in Figure 2.

anion for **2-CIP**. This indicates that the charge on the carbanion center is not extensively delocalized, a result that is consistent with a somewhat pyramidal carbanion center on the six-membered chair dithiane ring. It is interesting to compare the extent of charge delocalization in **1L** and **2L** with that in the anions derived from diphenylmethane and *p*-benzylbiphenyl, whose acidities in THF are comparable to **1** and **2**, respectively. While **1L** is exclusively contact ion pair, the lithium salt of diphenylmethane has been shown to contain approximately an equal amount of contact and solvent separated ion pairs.[21] Similarly, **2L** is mainly CIP while the UV–vis spectrum for the lithium salt of *p*-benzylbiphenyl indicates the presence of only SSIP and no detectable CIP. (The λ_{max} for the latter is 560.5 nm as compared to 529 nm for its cesium counterpart.) Thus, the dithiane group stabilizes the carbanion by about as much as a phenyl ring. However, different mechanisms appear to be involved in charge stabilization by the dithiane moiety in **1L** or **2L** and by the phenyl ring in the anions from diphenylmethane or benzylbiphenyl, since delocalization in the latter case promotes the formation of SSIP more than in the former (**1L** and **2L**). It may be that this difference in charge stabilization with respect to the formation of SSIP vs CIP could serve as an operational criterion to distinguish polarization from a delocalization mechanism.[22]

For carbon acids that form highly delocalized carbanions, the pK_a values normally show good agreement among different solvents[23] (such as DMSO, THF, and CHA) and between different counterions (such as lithium and cesium).[20] Indeed, our measured lithium pK_a value of 29.3 for **2** with respect to its SSIP is, within experimental error, in good agreement with the pK_a value in DMSO (29.1)[23b] and the pK_a value for the corresponding cesium salt in THF (29.1). This agreement validates our new method for determining the equilibrium lithium ion pair acidities of **1** and **2**.

Table 7. Dissociation Constant Ratios for Some Lithium and Cesium Salts in THF at 25 °C

compd	$K_{d(RLi)}/K_{d(RCs)}$ (M)
1L/1C	58.1
2L/2C for **2L-CIP**	34.2
2L/2C for **2L-SSIP**	390[a]
fluorene	282[b]

[a] This ratio is calculated from $K_{d(RLi)}/K_{d(RCs)}$ and [CIP]/[SSIP]. [b] From Table 5.

Table 8. Rate Constants[a] of Reactions of 1 and 2 with BDPMLi ((Biphenylyldiphenylmethyl)lithium) and BBLi ((Biphenylylbenzyl)lithium) in THF at 25 °C

	BDPMLi			BBLi		
compd	10^4k_{obs} (s^{-1})	10_2[**RH**] (M)	10^3k_2 ($M^{-1}\ s^{-1}$)	10^4k_{obs} (s^{-1})	10_2[**RH**] (M)	10^2k_2 ($M^{-1}\ s^{-1}$)
1	1.92	3.84	5.00	4.54	1.54	2.95
	3.84	7.93	4.84	6.76	2.06	3.28
				5.67	1.94	2.92
				4.72	1.58	2.99
				4.20	1.37	3.07
mean			4.92			3.04
2	3.39	1.31	25.9	4.97	0.217	22.9
	4.28	1.56	27.4	13.3	0.611	21.8
				37.0	1.59	23.2
mean			26.7			22.6

[a] Since a large number of data points (typically a few hundred to a few thousand data points) was acquired, the standard deviation of each k_{obs} from the nonlinear least-squares curve fit was found to be on the order of 10^{-6} to 10^{-7} s^{-1}.

From eq 8 and Table 5, the dissociation constant ratios $K_{d(RLi)}/K_{d(RCs)}$ can be calculated for both **1** and **2**. These ratios are listed in Table 7 along with the corresponding ratio for fluorenyllithium and fluorenylcesium. Since K_5 for compound **2** in Table 5 was measured for the equilibrium between **2L-CIP** and cesium tetraphenylborate, the above ratio for **2** refers to the relative extent of dissociation of **2L-CIP** to **2C** (the cesium salt of **2**). Since the dissociation constants for **2L-CIP** and **2L-SSIP** were measured independently, as discussed above (Dissociation of 2L to Free Ion), the corresponding dissociation constant for **2C** to the free ions can be calculated to be 3.6×10^{-8} M. The magnitude is comparable to that for fluorenylcesium at 25 °C in THF ($K_d = 1.38 \times 10^{-8}$ M).

Brønsted *β* Values. The slow reaction between **1** or **2** and the indicator salts, such as (*p*-biphenylylbenzyl)lithium (**BBLi**) or (biphenylyldiphenylmethyl)lithium (**BDPMLi**), allowed us to study the kinetic acidities of **1** and **2**. The experimental details of these studies are given in the Experimental Section. The observed first-order rate constants and the calculated second-order rate constants are listed in Table 8. The calculations of Brønsted *β* values require known values of lithium equilibrium acidities of *p*-benzylbiphenyl and biphenylyldiphenylmethane. Unfortunately, these values have not been measured in THF because of the same slow establishment of equilibrium found for **1** and **2**. These two compounds do form delocalized carbanions (exclusively SSIP observed with substantial dissociation to free ion, but no CIP) and, as discussed earlier, should have lithium pK_a values close to their cesium pK_a values (31.7 for *p*-benzylbiphenyl and 30.07 for biphenylyldiphenylmethane).[18c] On the basis of these equilibrium pK_a values, the calculated Brønsted *β* values for **1** and **2** are 0.48 and 0.57, respectively. The uncertainty associated with these values is somewhat greater due to the error

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[21] Buncel, E.; Menon, B. C.; Colpa, J. P. *Can. J. Chem.* **1979**, *57*, 999.

[22] For discussion on the mechanism of anionic stabilization by sulfur, see: (a) Williams, J. E., Jr.; Streitwieser, A., Jr. *J. Am. Chem. Soc.* **1975**, *97*, 191. (b) Epiotis, N. D.; Yates, R. L.; Bernardi, F.; Wolfe, S. *J. Am. Chem. Soc.* **1976**, *98*, 7498. (c) Borden, W. T.; Davidson, E. R.; Andersen, N. H.; Denniston, A. D.; Epiotis, N. D. *J. Am. Chem. Soc.* **1978**, *100*, 1604. (d) Schleyer, P. v. R.; Clark, T.; Kos, A. J.; Spitznagel, G. W.; Rohde, C.; Arad, D.; Houk, K. N.; Rondan, N. G. *J. Am. Chem. Soc.* **1984**, *106*, 6467. (e) Bernardi, F.; Mangini, A.; Tonachini, G.; Vivarelli, P. *J. Chem. Soc., Perkin Trans. 2* **1985**, 111. Boche, G.; Lohrenz, J. C. W.; Ciolowski, J.; Koch, W. In *The Chemistry of Sulphur-Containing Functional Groups*; Patai, S., Rappoport, Z., Eds.; Wiley: New York, 1993; Supplement 5, pp 340–362.

[23] (a) Streitwieser, A., Jr.; Juaristi, E.; Nebenzahl, L. L. In *Comprehensive Carbanion Chemistry*; Buncel, E., Durst, T., Eds.; Elsevier: New York, 1980. (b) Bordwell, F. G.; Drucker, G. E.; Andersen, N. H.; Denniston, A. D. *J. Am. Chem. Soc.* **1986**, *108*, 7310. (c) Reference 9.

associated with the measured rate constants and to the direct use of cesium equilibrium pK_a values in these calculations. Nevertheless, the results are of qualitative utility. Both values are near 0.5 and suggest that the transition state for the proton transfer process in these kinetic studies is rather symmetric. This result is reasonable since both *p*-benzylbiphenyl and biphenylyldiphenylmethane are comparable in acidity to **1** and **2**.

Conclusions

We have described a new approach to determining the equilibrium lithium ion pair acidities of 2-phenyl-1,3-dithiane (**1**) and 2-*p*-biphenylyl-1,3-dithiane (**2**), which are otherwise difficult to obtain. The lithium salt of **1** is exclusively contact ion pair, while that of **2** consists of a mixture of contact ion pair, solvent-separated ion pair, and free ions. The dissociation of solvent-separated ion pair to free ion is especially significant at dilute concentrations. Such dissociation constants were measured in THF and found to be comparable in magnitude to those measured previously for fluorenyllithium in THF. The enthalpy and entropy change for the equilibrium SSIP $\rightleftharpoons$ CIP were also measured and indicate that SSIP is preferred at lower temperatures. The acid strength-enhancing moiety, the dithianyl group, stabilizes the carbanion charge, but is inferior to a phenyl ring in promoting the formation of SSIP. This difference might serve as an operational criterion for distinguishing polarization from delocalization mechanisms of charge stabilization.

Experimental Section

General. Glassware was dried in an oven at 160 °C for at least 4 h prior to use. All measurements of transmetalation equilibrium constants and rate constants were made in a glovebox filled with argon. UV−vis spectra were recorded on a Shimadzu UV-2101PC spectrophotometer equipped with fiber optics cables connected to a thermostated cell holder built into the floor of the glovebox. Constant temperature was achieved by using a Neslab Model RTE-8DD temperature controller for room temperature and higher and by using a Neslab Model ULT-8ODD for lower temperatures, using 95% ethanol as the coolant. Proton nuclear magnetic resonance (^{1}H-NMR) spectra were conducted on UCB-250 (a superconducting 250 MHz instrument), with Bruker AM-400 and Bruker AM-500 spectrometers operating in the FT mode. Gas−liquid chromatography (GLC) was performed on a Hewlett-Packard 5880 gas chromatograph using a 6 ft × 18 in. OV-101 column. Melting points (Pyrex capillary) were determined on a Buchi melting point apparatus and are uncorrected.

Materials. 2-Phenyl-1,3-dithiane (**1**), 2-*p*-biphenylyl-1,3-dithiane (**2**), and indicators used were available from our previous studies.[9] The purity of the compounds was monitored by combinations of ^{1}H-NMR, GLC, mp, and elemental analysis. (Biphenylylbenzyl)lithium (**BBLi**) was prepared in the glovebox by treating *p*-benzylbiphenyl with excess lithium metal in THF at room temperature. After 2 days, the purple solution of the anion was decanted away from excess lithium. The UV−vis spectrum of **BBLi** shows a maximum absorbance at 560.5 nm. (Biphenylyldiphenylmethyl)lithium (**BDPMLi**) was prepared similarly. Its UV−vis spectrum shows a maximum absorbance at 602.5 nm. Ethyllithium was prepared by Todd McDermott.

Lithium tetraphenylborate was obtained from Aldrich Chemical Co. and dried in a Schlenk flask at 150 °C under high vacuum (15 μm) for 12 h immediately prior to use. Ethyllithium was synthesized according to a literature procedure.[24] The crude product was sublimed under vacuum (80 °C at 0.05 mm Hg) and stored in a freezer at −30 °C in the glove box: ^{1}H-NMR (C_6D_6) δ 1.26 ppm (t, 3H), −0.94 ppm (q, 2H).

UV−vis Spectra. The cesium salts of 2-phenyl-1,3-dithiane (**1C**) and 2-*p*-biphenylyl-1,3-dithiane (**2C**) were prepared by deprotonating **1** and **2** with cumylcesium. The corresponding lithium salts (**1L** and **2L**) were prepared by adding excess lithium tetraphenylborate to solutions of **1C** and **2C**. The exchange of cations is rapid, and the recorded spectra of **1L** and **2L** were stable and identical to those obtained individually by deprotonating **1** and **2** with ethyllithium. The extinction coefficients of these anions at their λ_{max}'s are tabulated in Table 1. The data for **1C** and **2C** obtained previously are included for comparisons. The extinction coefficients for **1L** and **2L** were determined by titrating a known amount of **1** or **2** with ethyllithium in the presence of lithium tetraphenylborate in order to suppress free ion formation. The reactions of ethyllithium with **1** and **2** were instantaneous at room temperature.

Purification of THF. Fisher Scientific reagent grade THF was distilled from lithium aluminum hydride or from sodium−benzophenone into an oven-dried Schlenk flask containing a magnetic stirring bar. The flask was attached to a vacuum line, and the contents were degassed by at least three freeze−pump−thaw cycles. It was then taken into the glovebox, and several drops of sodium−potassium alloy were added. The solution was stirred in the glovebox until a blue color persisted. The flask was then taken out of the glovebox and attached to the vacuum line, and the contents were degassed by three freeze−pump−thaw cycles. During the degassing process, the blue color of THF-solvated electron sometimes disappeared due to freezing. In such cases, the solution was stirred after degassing until a blue color reappeared, and THF was vacuum transferred to a long neck receiving flask which had been oven- and flame-dried. The contents were again degassed and taken into the glovebox, and approximately 20 g of 3 Å molecule sieves were added for every 250 mL of THF solvent. The solvent was allowed to stand over the sieves for at least 2 days in a quartz flask prior to use. The water content is estimated to be on the order of 10^{-6} M.

Acidity Measurements. The methods of acidity measurements used here involve some modifications of our double indicator method described previously.[9] A solution of **1C** (or **2C**) in THF was prepared in a 0.1 cm path length quartz cell by deprotonating **1** (or **2**) with cumylcesium. A known amount of lithium tetraphenylborate was weighed into a quartz microbeaker, and the beaker was dropped into the cell. The solution was shaken until all of the lithium tetraphenylborate dissolved. The mixed spectra of **1C** and **1L** (or **2C** and **2L**) were then recorded and deconvoluted by a linear least-squares method (described below) based on the inputs of the standard spectra of **1C** and **1L** (or **2C** and **2L**). An example of such a deconvoluted spectrum is shown in Figure 4. The equilibrium constants for the following reactions were then calculated:

$$\mathbf{1C} + LiBPh_4 \rightleftharpoons \mathbf{1L} + CsBPh_4$$

$$\mathbf{2C} + LiBPh_4 \rightleftharpoons \mathbf{2L} + CsBPh_4$$

These equilibrium constants, combined with the dissociation constants of cesium and lithium tetraphenylborate ($K_{d(LiBPh4)}$ and $K_{d(CsBPh4)}$ in Table 5, Results and Discussion), allow the calculations of the ratios of the dissociation constants, $K_{d(1L)}/K_{d(1C)}$ or $K_{d(2L)}/K_{d(2C)}$. From these ratios and the corresponding ratios of cesium and lithium salts of fluorene indicator, the lithium pK_a values of **1** and **2** can then be derived.

Since these measurements were conducted in the presence of lithium tetraphenylborate, we used a standard spectrum of **2L** recorded in the presence of lithium tetraphenylborate for the deconvolution. The use of lithium tetraphenylborate eliminated the possible spectral interference from the free ion in the deconvolution process by suppressing the free ion formation.

Deconvolution of Mixed Spectra. The standard UV−vis spectra of **1C** and **1L** were obtained experimentally by

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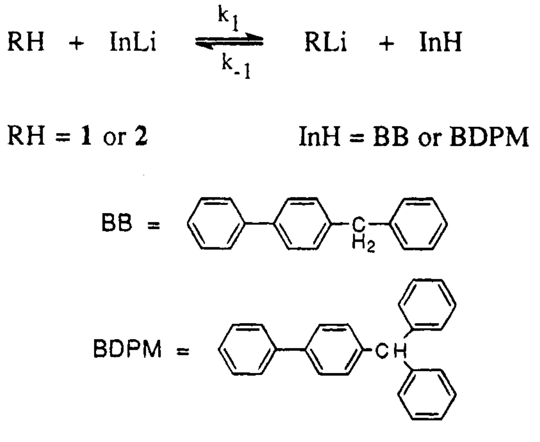
[24] Masamune, S.; Choy, W. *Aldrichim. Acta* **1982**, *15*(3), 47.

Scheme 2. Reactions of 1 and 2 with Indicator Anions with Lithium as Counterion



scanning solutions of **1C** and **1L** separately. Their concentrations were adjusted so as to give UV–vis spectra with maximum absorbance of about 1.0. The spectra were stored in the computer and converted to ASCII text files containing typically a few hundred to 1000 pairs of X–Y data points. These data points represent the absorbance (*Y* axis) at each corresponding wavelength (*X* axis). Thus, the absorbance of **1C** and **1L** can be described by arrays $Y_C(1)$, $Y_C(2)$, ..., $Y_C(n)$, and $Y_L(1)$, $Y_L(2)$, ..., $Y_L(n)$, respectively, where *n* is the number of data points.

The mixed spectrum of **1C** and **1L** was obtained in the acidity measurement when the two components were allowed to equilibrate. Most neutral compounds have negligible absorption in the wavelength range where **1C** and **1L** absorb dominantly. The absorbance of the mixed spectrum can be represented by the array $Z(1)$, $Z(2)$, ..., $Z(n)$. A common way to calculate the relative contributions of the two components is by solving two linear equations based on two chosen points at which the molar absorptivities of both components are known. However, we used a linear least-squares method which should give better deconvolution results since the entire range of absorbances was used. The linear least-squares method in this case involves minimizing the sum of the squared difference between the absorbance of the mixed spectrum and that of the calculated spectrum. Thus, the sum of squared difference, SD, is described by

$$\mathrm{SD} = \sum[Z(i) - (F_1Y_C(i) + F_2Y_L(i)]^2 \quad i = 1,2, ..., n$$

where F_1 and F_2 are the multiplying factors for the relative contributions of the two standard spectra.

To minimize SD, the partial derivatives of SD with F_1 and F_2 are set to zero:

$$\sum Y_C(i)[Z(i) - (F_1Y_C(i) + F_2Y_L(i)] = 0$$
$$\text{and} \sum Y_L(i)[Z(i) - (F_1Y_C(i) + F_2Y_L(i)] = 0$$

A simple computer program was written to solve for the factors F_1 and F_2 in the above two equations. The absorbance of the calculated spectrum ($F_1Y_C(i) + F_2Y_L(i)$) was then graphically displayed along with the mixed spectrum for comparisons. Note the excellent agreement between the two in Figure 4.

The above treatment used **1C** and **1L** as the examples. The treatment applies as well to other systems, such as **2C** and **2L**, and can be readily extended to more than two components.

ΔH° and ΔS° Measurements for Equilibrium SSIP ⇌ CIP for 2L. A solution of **2L** in THF was prepared by deprotonating **2** with ethyllithium, and excess lithium tetraphenylborate was added to suppress free ion. The maximum absorbance (λ_{max}) for CIP and SSIP of **2L** is 388 nm and 485 nm, respectively, with the ratio of CIP/SSIP = 11.4 at 25 °C. This ratio is independent of **2L** concentration *in the presence of excess lithium tetraphenylborate*, indicating the effective suppression of free ion formation in such solutions. The equilibrium constant for SSIP ⇌ CIP was measured at several other temperatures ranging from 25 to −20 °C, and the results are summarized in Table 4. By plotting ln *K* as a function of 1/*T*, $\Delta H° = 1.43$ kcal mol^{-1} and $\Delta S° = 9.6$ eu were obtained from the Van't Hoff plot (Figure 2).

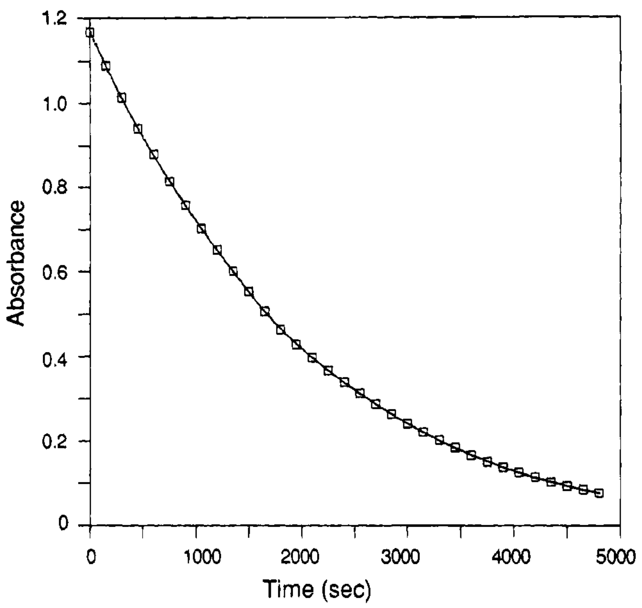


Figure 5. Absorbance of **BBLi** ((biphenylylbenzyl)lithium) at 562.5 nm as a function of time for the reaction of **1** with **BBLi**. The squares are experimental data points; for clarity only 33 points out of 3200 points are shown. The solid line is the theoretically fitted curve to the first order exponential decay.

Dissociation Constants of 2L to Free Ion. Unlike **1L**, which exists exclusively as contact ion pair in THF, **2L** consists of a mixture of CIP, SSIP, and free ion. Because of the small difference between the λ_{max} of SSIP and that of the free ion, their spectra overlap to form a broad peak whose λ_{max} changes with **2L** concentration. Obviously, the λ_{max} of the mixed spectra depends on the relative contributions of the two components (SSIP and free ion) which should be concentration dependent. The relative amount of the free ion increases as the solution is diluted, since the free ion formation is a dissociation process. Figure 1 illustrates the shift of λ_{max} (around 490 nm) as the concentration of **2L** is changed. Note the decrease in the relative ratio of the maximum absorbance for the CIP peak at 388 nm to the peak (at about 490 nm) due to the mixture of SSIP and the free ion as the concentration is lowered, which is consistent with the greater free ion. contribution. Addition of excess lithium tetraphenylborate to different concentrations of **2L** significantly reduced the free carbanion concentration, and λ_{max} for SSIP was at 485 nm while that for CIP was unchanged. Furthermore, the same ratio of [CIP]/[SSIP] was obtained for different concentrations of **2L**, and the resulting spectra were insensitive to the addition of more lithium tetraphenylborate. The spectra of CIP plus SSIP (with excess lithium tetraphenylborate) were subtracted from the mixed spectra of CIP, SSIP, and the free ion to obtain the free ion peak at λ_{max} 506 nm. The dissociation constants can thus be calculated assuming identical extinction coefficients for CIP, SSIP, and the free ion. The results are listed in Table 2. From the thermodynamic cycle in Scheme 1, $K_{d(SSIP)} = 11.40K_{d(CIP)} = 1.39 \times 10^{-5}$ M (± 0.09).

The CIP peak (λ_{max} 388 nm) and the SSIP peak (λ_{max} 485 nm) have an overlap region regardless of whether an excess amount of lithium tetraphenylborate is added or not. The absorbance readings at the λ_{max}'s of both peaks, therefore, should not be used to calculate the relative ratio of CIP/SSIP unless it can be shown that the two peaks do not contribute significantly to each other at 388 and 485 nm. This is obviously difficult to show experimentally, but a linear curve fitting process was used for this purpose. The mixed spectra of CIP and SSIP were deconvoluted using two combined Gaussian functions (three parameters per function: center, amplitude, and width), whose linear combinations describe very well the experimental spectra.[25] The results show that the CIP peak has negligible contribution to the absorbance of the SSIP peak at λ_{max} and vice versa. Thus, the absorbance readings at both λ_{max} can be directly used in the calculation of CIP/SSIP ratio.

Kinetic Studies of 1 and 2. While transmetalation reactions are generally fast with cesium as counterion, the rates of equilibration with lithium counterion are extremely slow for compounds whose pK_a's are above about 25. These compounds include the dithianes **1** and **2**, which were found to react with indicator anions so slowly that it would have taken days for them to reach equilibrium. Therefore, it is difficult to measure the equilibrium lithium ion pair acidities of **1** and **2** with the conventional methods used previously for the cesium ion pair acidities. These slow transmetalation reactions, on the other hand, allow the convenient measurements of their kinetic acidities. **1** and **2** were thus allowed to react with two different lithium indicator bases, (biphenylylbenzyl)lithium (**BBLi**) and (biphenylyldiphenylmethyl)lithium (**BDPMLi**), as shown in Scheme 2

To measure the forward rate constant k_1, a large excess of neutral RH (**1** or **2**, at least 30-fold excess over **BBLi** or **BDPMLi**) was used to ensure that the reactions follow pseudo-first-order kinetics. The reactions were conveniently followed by recording the absorbance of **BBLi** at its λ_{max} of 562.5 nm or the absorbance of **BDPMLi** at 602.5 nm. These absorbance readings were not interfered with by the spectra of **1L** or **2L** and were recorded at desired time intervals for more than 3 half-lives in most of the runs. The resulting exponential decay data (usually several hundred to a few thousand data points) were analyzed by a nonlinear curve fitting to the normal first-order kinetic equation. An illustrative plot of the absorbance vs time is shown in Figure 5. The second-order rate constants were then calculated and are tabulated in Table 8.

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(25) Gronert, S.; Streitwieser, A., Jr. *J. Am. Chem. Soc.* **1988**, *110*, 2836–42.