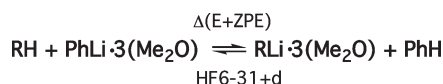# Ab Initio Modeling of Organolithium Equilibria

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$$\text{RH} + \text{PhLi·3(Me}_2\text{O)} \underset{\text{HF6-31+d}}{\overset{\Delta(\text{E+ZPE})}{\rightleftharpoons}} \text{RLi·3(Me}_2\text{O)} + \text{PhH}$$

Experimental ion pair p$K$'s of monomeric contact ion pair lithium salts in THF from our previous studies give good correlations with ab initio calculations at the Hartree−Fock 6-31+g(d) level. PCM methods were found to be inadequate in nonpolar organic solvents, and dielectric solvation was not used in the correlations. Specific coordination of two or three ether solvent molecules with lithium was found to be satisfactory. These correlations include carboxamides, amines, dithianes, sulfones, and sulfoxides, as well as some ketones, $\beta$-diketones, and the lithium salts of dianions.
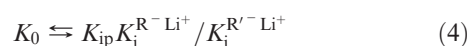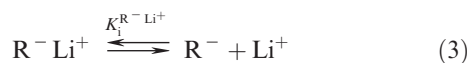
## Introduction

Over a period of years we have measured a number of lithium ion pair acidities as defined by the lithium exchange equilibrium, eq 1. Our initial measurements were in cyclohexylamine, but for the past two decades these measurements were made in THF, a more important solvent for synthetic chemistry. Particular attention was paid to aggregation, and the measurements were generally shown to apply to the monomers or were corrected for the presence of aggregates.

$$\text{RH} + \text{R}'^{-}\text{Li}^{+} \overset{K_{\text{ip}}}{\rightleftharpoons} \text{R}^{-}\text{Li}^{+} + \text{R}'\text{H} \tag{1}$$

Equation 1 shows an apparent relationship to eq 2, used by Bordwell[1] to determine the acidities of hundreds of compounds in DMSO, but there is an important difference between them, beyond the distinction of ion pairs versus ions.

$$\text{RH} + \text{R}'^{-} \overset{K_{0}}{\rightleftharpoons} \text{R}^{-} + \text{R}'\text{H} \tag{2}$$
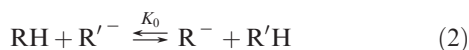
In both cases one of the components is treated as an indicator and the equilibria define the $\Delta pK_a$ between it and the substrate. The equations are related by use of the ion dissociation equilibria of the ion pairs, eqs 3 and 4:

$$\text{R}^{-}\text{Li}^{+} \overset{K_{\text{i}}^{\text{R}^{-}\text{Li}^{+}}}{\rightleftharpoons} \text{R}^{-} + \text{Li}^{+} \tag{3}$$

$$K_{0} \rightleftharpoons K_{\text{ip}} K_{\text{i}}^{\text{R}^{-}\text{Li}^{+}} / K_{\text{i}}^{\text{R}'^{-}\text{Li}^{+}} \tag{4}$$

A number of $K_{\text{i}}^{\text{R}^{-}\text{Li}^{+}}$ values have been determined[2] that could in principle relate $K_{\text{ip}}$ and $K_{0}$, except for an essential distinction. Bordwell[3] included as indicators several relatively acidic compounds (p$K$ < 12) whose p$K_a$'s were determined in DMSO by conductometric or potentiometric methods.[4−9] Thus, although Bordwell only determined $\Delta pK$ values, by including compounds of known p$K_a$ all of his acidities are therefore p$K_a$ values in DMSO.

Only a few p$K_a$'s have been determined in THF and then only for such strong acids as perchloric acid.[10] No equilibrium measurements of the type of eqs 1 or 2 are known with these strong acids in THF; thus, the p$K$'s measured in THF *are not* p$K_a$'s. That is, the numbers often referred to in the literature as p$K_a$ in THF are not in fact p$K_a$ values. Instead, one compound is chosen as an arbitrary standard, and the measured $\Delta pK$ values define a p$K$ *relative to this assumed standard*. We have used the solvent separated ion pair (SSIP) of fluorenyllithium as our reference standard and have assigned to it its p$K_a$ in DMSO, 22.90 (per hydrogen; the value given by Bordwell, 22.60, does not include the statistical factor of 2). Others in the literature have used different

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**TABLE 1. Comparison of CIP pK(Li,THF) with pK$_a$ in DMSO$^a$**

| compound | pK(Li,THF) | ref | compound | pK$_a$(DMSO) | ref |
|---|---|---|---|---|---|
| diphenylamine | 19.05 | 12 | diphenylamine | 25.0 | 13 |
| carbazole | 13.48 | 14 | carbazole | 19.9 | 15 |
| PhCH$_2$SO$_2$Ph | 19.2 | 16 | PhCH$_2$SO$_2$Ph | 23.4 | 3 |
| PhCH$_2$COCH$_2$Ph | 11.32 | 17 | PhCH$_2$COCH$_2$Ph | 18.7 | 18 |
| 6-Ph tetralone | 13.92 | 19 | tetralone | 24.7 | 20 |
| BiphCH$_2$CONMe$_2$$^b$ | 19.47 | 21 | PhCH$_2$CONMe$_2$ | 26.6 | 22 |
| BiphCH$_2$SOPh$^b$ | 19.8 | 16 | PhCH$_2$SOPh | 27.2 | 23 |
| adamantylC≡CH | 23.7 | 24 | PhC≡CH | 28.8 | 3, 25 |
| BiphCOCHMe$_2$$^b$ | 15.86 | 26 | PhCOCHMe$_2$ | 26.3 | 18 |
| BiphCOCH$_2$COCH$_3$$^b$ | 1.18 | 27 | PhCOCH$_2$COCH$_3$ | 14.2 | 18 |

$^a$All pK's are *not* statistically corrected. $^b$Biph = *p*-biphenylyl.

reference standards. In a future paper we will attempt to reconcile these disparate acidity numbers.

We symbolize our values as pK(Li,THF) or, when the specification of solvent is obvious, as pK(Li). For a number of SSIP lithium salts of highly delocalized carbanions of the fluorenyl and benzofluorenyl type, the $K_i$ values are relatively constant at $10^{-5}$ M.[2] For these compounds the pK(Li) closely parallels the pK$_a$ in DMSO as expected from eq 4.

The linear relationship between these scales, eq 5, was published previously;[11] the slope close to unity means that any other compound in this correlation could have been chosen equally well as a standard without significant change in the numerical values of pK(Li,THF).

$$pK(Li, THF) = -0.963 + 1.046pK_a(DMSO) \quad (5)$$

The lithium salts of localized carbanions and many similar salts with Li−O and Li−N bonds are contact ion pairs (CIP) in which the anions are more tightly bound to the lithium cation than the SSIP. In accordance with eq 4, the pK(Li, THF) of such compounds are generally *lower* than the corresponding pK$_a$ in DMSO. Some examples are summarized in Table 1. They include only four compounds measured in both solvents and for which a direct comparison is possible. Most of the remaining compounds, however, differ only by a phenyl substituent that should have only a minor effect on the pK. Our pK(Li,THF) measurements generally used substrate and indicator salts both of which have suitable UV−vis spectra and could be measured for greater accuracy

(the "double-indicator method"). Most of the DMSO substrates do not have such spectra and were measured by the so-called single-indicator technique. Nevertheless, for all of the compounds the differences between similar compounds in DMSO and THF are so great that the generalization is clearly established. The differences also vary substantially, from about 4 to 13 pK units. We were thus led to consider the correlation of the CIP pK(Li,THF) scale with ab initio computations.

**Methodology**

The computation of RH and RLi is straightforward for the gas phase, but in THF solutions the solvation energies cannot be neglected. We have used previously a polarized continuum model (PCM) in theoretical studies of some reaction mechanisms,[28] and Fu et al.[29] have used PCM in their theoretical treatment of cesium ion pair acidities. In recent calibration studies, however, we found that PCM is ineffective for solvation energies of ordinary organic compounds in nonpolar solvents such as THF and ethyl ether. A typical example comparing calculated and experimental solvation energies in Figure 1 shows an almost random scattering of points. There are many more solvation energies available for ethyl ether than for THF, but for several compounds measured in both solvents, the experimental values are quite similar.

The solvation energies are relatively small and cover a small range. Cramer and Truhlar[30] have also reported deficiencies of various PCM. They report errors comparable to the values of the solvation energies and even suggest that the use of a single average value for all compounds is almost as successful! They have developed a new method, SM8, a multidimensional and highly parametrized method that gives much better results.[30,31] This method, however, still has significant deficiencies and, moreover, is not parametrized for organolithium compounds. Accordingly, we have neglected dielectric solvation in the present study. By casting the results in the isodesmic-type eq 6, however, we anticipate a substantial cancellation of solvation terms. We do include, as we[28,32,33] and others[34−36] have done, specific coordination of solvent molecules to the lithium cation,

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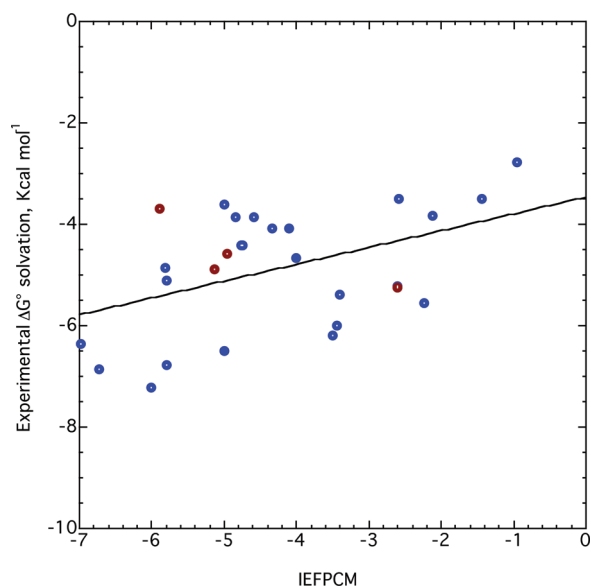(23) Bordwell, F. G.; Branca, J. C.; Johnson, C. R.; Vanier, N. R. *J. Org. Chem.* **1980**, *45*, 3884–3889.
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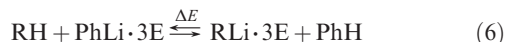(26) Abbotto, A.; Leung, S. S.-W.; Streitwieser, A.; Kilway, K. V. *J. Am. Chem. Soc.* **1998**, *120*, 10807–10813.
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**FIGURE 1.** Experimental $\Delta G°$(solvation) in ethyl ether (blue) and THF (red) compared to IEFPCM computations. The regression line shown is $-3.46 \pm 0.61 + (0.311 \pm 0.134)x$; $R^2 = 0.21$. The data are given in Table S1 in Supporting Information.

but because of the large size of some of our systems, we use dimethyl ether (E) as our model for THF. With these additions our computational model for organolithium equilibria is eq 6, in which lithium is treated as four-coordinate and the equilibria are considered relative to phenyllithium as the arbitrary reference.

$$RH + PhLi \cdot 3E \underset{}{\overset{\Delta E}{\rightleftharpoons}} RLi \cdot 3E + PhH \qquad (6)$$

Calculations were done at the HF level with the 6-31+G* basis set . The diffuse function was added because we are dealing with salts of anions and to reduce the magnitude of any basis set superposition errors (BSSE). Frequencies were calculated with the same basis set, and $\Delta E$ was taken as the computed energy plus the unscaled ZPE using Gaussian03[37] or Gaussian09.[38] All of the compounds in this study are minima on the potential energy surface (PES) with all frequencies real. Ideally, one should use the free energy for direct comparison with equilibrium constants at 25 °C, but the three ethers coordinated with lithium produce a number of low frequency vibrations whose anharmonicity introduce substantial error into the calculation of thermodynamic quantities (especially the entropy) in which all vibrations are treated as harmonic. For these lithium compounds coordinated with ethers, typically half or more of the vibrational entropy is associated with vibrations less than 100 cm$^{-1}$.

Since the coordination of ethers with lithium is weaker than normal bonds, one might expect structures with alternative arrangements of the coordinating ethers. We did not make a complete search for all such alternate structures, but in a number of cases we did do additional optimizations with different starting structures. Most such studies yielded the same final structure, but alternative conformations were found in several cases. In some such cases the structures differ but little and have similar energies. In some other cases, however, the structures and energies differ substantially. An example is shown in Figure 2. The two structures do not differ much but B is 0.49 kcal mol$^{-1}$ lower in energy. The principal difference is the Li−O bond that is shorter in B (1.776 vs 1.802 Å) and the C−O−Li angle that is more linear (156.3° vs 135.7°). Both structures are minima on the

PES (all frequencies real), and the three Li−Me$_2$O bonds are similar. The dihedral angle between the phenyl group and the enolate double bond is slightly more conjugating (16.5° vs 20.1°) in B.
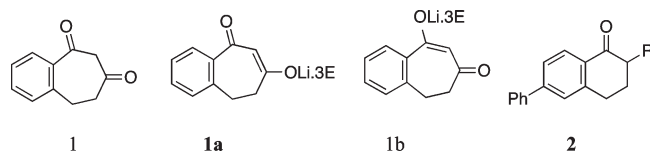
In all cases in which alternative conformations were found, only the lowest energy conformation was considered. A number of compounds, particularly lithium enolates, exist as *Z* and *E* isomers. Here also, we used only the isomer of lowest energy.

In principle, a Boltzmann distribution of all such conformations should be used, but this additional step was not taken. The neglect of dielectric solvation probably involves errors of at least comparable magnitude. The coordinates, ZPE, and thermal corrections of all of the structures computed in this work are summarized in Supporting Information. The energies (including the ZPE) and experimental p$K$(Li) values used are summarized in Table 2. With one exception, these lithium compounds are known to be CIP in THF at room temperature. 2-Phenyl-1,3-dithianyllithium is known to be completely a CIP,[39−41] but the corresponding biphenylyl compound contains a small amount of SSIP that adds a small correction to its p$K$(Li).[42]

## Results and Discussion

A plot of the experimental p$K$(Li) values against the computed $\Delta E$ values is shown in Figure 3. The correlation is excellent with only the tetralone compounds deviating seriously, probably because of the neglect of dielectric solvation. The excellent fit of most of the other compounds indicates that in the form of eq 6 solvation energies cancel to a large extent.

The single $\beta$-diketone included is the benztropadione compound, **1**, in which the carbonyl groups are directed away from each other (W-shaped) and the oxygens are too far apart to mutually coordinate lithium in the enolate ion. That is, the isomeric lithium enolates are distinctly different with **1b** slightly more stable than **1a**. The point shown in Figure 3 is that for **1b** only. It is noteworthy that it deviates slightly from the correlation line in the same direction as the tetralones, **2**, which have the same type of bicyclic structure and could therefore have similar solvation properties.



Among the benzene derivatives, only 1,2,4,5-tetrafluorobenzene was measured directly for its p$K$(Li); pentafluorobenzene and benzene itself were estimated by analogy with the corresponding cesium compounds.[43] These compounds, and especially benzene itself, fit the general correlation. The several carboxamides correlate well, as do the two amines, diphenylamine and carbazole. With the carboxamides, even subtle changes, such as NEt$_2$ instead of NMe$_2$, are mirrored by the computations. Overall, except for the bicyclic compounds, the p$K$(Li) are predicted to an average of ±0.7 p$K$ units.

For these same compounds, calculations were made at the MP2(6-31+d) level using the HF geometries. The results

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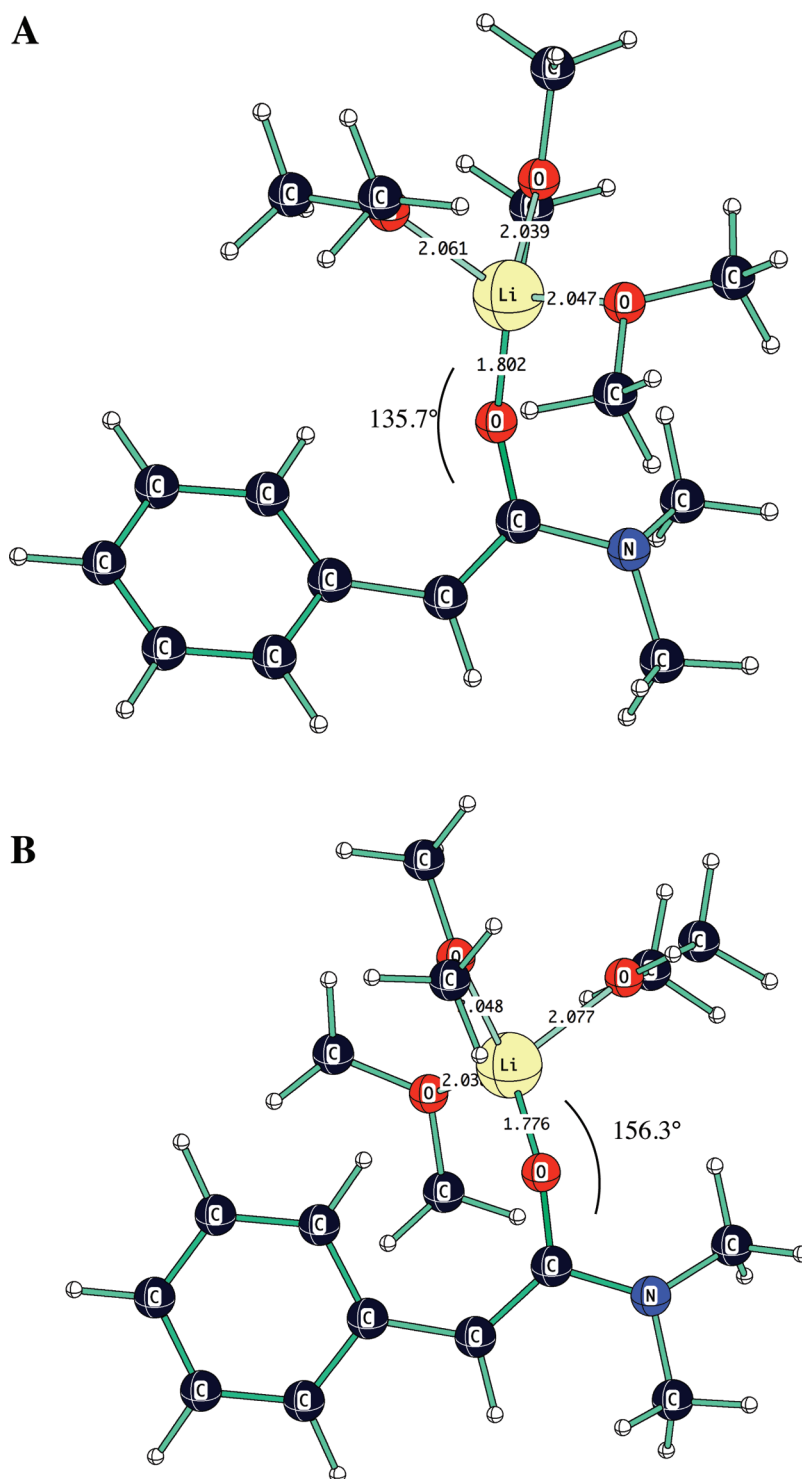(39) Abatjoglou, A. G.; Eliel, E. L.; Kuyper, L. F. *J. Am. Chem. Soc.* **1977**, *99*, 8262–8269.
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**A**



**B**



**FIGURE 2.** Two structures for the *Z* lithium enolate of phenyl-(*N*,*N*-dimethylacetamide) coordinated to three dimethyl ethers. Structure B is lower in energy than A by 0.49 kcal mol$^{-1}$.

(Figure S1, Supporting Information) show more scatter than Figure 3 and a distinctly inferior correlation. $\Delta G°$ for eq 6 was also compared with p$K$(Li) as shown in Figure S2 (Supporting Information), using $G_{\text{thermal}}$ as calculated by Gaussian. The plot is similar to Figure 3 but with somewhat more scatter and an inferior correlation ($R^2 = 0.92$) compared with E + ZPE. The results in Figure 3 are at 0 K. We included the thermal correction as calculated by Gaussian by plotting p$K$(Li) against

$E + E_{\text{thermal}}$. The resulting plot, shown as Figure S3 (Supporting Information), is almost exactly the same as Figure 3 with $R^2 = 0.982$. The largest contributions are those associated with the Li·3E group, and these tend to cancel in eq 6.

We next inquired whether it is necessary to use three solvents coordinated to lithium. Determination of the number of coordinated solvent molecules in ethereal solutions is not simple, and in many studies a four-coordinate lithium is assumed. At low

**TABLE 2.** Experimental p$K$(Li) and computed E + ZPE (HF, 6-31+G*) for RH, RLi·2E, and RLi·3E (E = dimethyl ether)

| | | | E + ZPE, au | | |
|---|---|---|---|---|---|
| RH | p$K$(Li,THF) | ref | RH | RLi·2E | RLi·3E |
| PhH | 39.5 | 43 | −230.603532 | −545.464582 | −699.457818 |
| PhDithiane | 29.4 | 42 | −1180.492291 | −1495.378592 | −1649.371381 |
| BiphDithiane | 28.2 | 42 | −1409.962082 | −1724.848951 | −1878.841908 |
| AdaC≡CH[a] | 23.7 | 24 | −463.435217 | −778.338293 | −932.333406 |
| C$_6$H$_2$F$_4$ | 23.1 | 43 | −626.038467 | −940.942741 | −1094.937831 |
| Ph$_2$CHCONEt$_2$ | 22.02 | 21 | −822.830080 | −1137.738701 | −1291.727560 |
| PhBnOxazole | 21.55 | 21 | −743.673246 | −1058.577929 | −1212.563628 |
| C$_6$HF$_5$ | 21.5 | 43 | −724.886046 | −1039.793217 | −1193.788347 |
| BiphCOCH(Li$_2$E)COCH$_3$ | 21.9 | 27 | −1078.508623 | −1393.405178 | |
| Ph$_2$CHCON pyrrolidine | 21.11 | 21 | −821.692623 | −1136.604967 | −1290.592238 |
| Ph$_2$CHCONMe$_2$ | 20.78 | 21 | −744.817269 | −1059.730977 | −1213.719510 |
| BiphCH$_2$CONEt$_2$ | 20.36 | 21 | −822.841858 | −1137.755202 | −1291.742580 |
| BiphMeSOPh | 20.1 | 16 | −1200.852909 | −1515.765485 | −1669.754646 |
| BiphDMA | 19.77 | 21 | −744.829043 | −1059.743706 | −1213.732335 |
| BnSO$_2$Ph | 19.5 | 16 | −1046.228017 | −1361.142352 | −1515.133247 |
| Ph$_2$NH | 19.05 | 12 | −515.074971 | −829.991309 | −983.982628 |
| BiphMeSO$_2$Ph | 18.8 | 16 | −1275.697768 | −1590.612493 | −1744.603822 |
| BnCOCH(Li$_2$E)COPh | 17.7 | 27 | −1078.503017 | −1393.413103 | |
| PhIBP[b] | 15.86 | 26 | −689.809802 | −1004.726207 | −1158.720829 |
| PhSO$_2$IBP[b] | 14.69 | 44 | −1236.959153 | −1551.878150 | −1705.873146 |
| 6-Ph tetralone | 14.22 | 19 | −688.684915 | −1003.598025 | −1157.593801 |
| 2-Bn,6-Ph tetralone | 13.96 | 45 | −957.154680 | −1272.071310 | −1426.064936 |
| Cbz[c] | 13.48 | 14 | −513.956369 | −828.883041 | −982.876720 |
| PhCHX[d] | 12.69 | 46 | −537.205752 | −852.136173 | −1006.126527 |
| BiphCHX[d] | 12.31 | 47 | −766.675490 | −1081.606391 | −1235.597038 |
| Bn$_2$CO | 11.62 | 17 | −650.805211 | −965.733873 | −1119.725277 |
| 2,6-diPh tetralone | 11.14 | 48 | −918.146297 | −1233.072857 | −1387.062542 |
| benzcycloheptadione[e] | 4.82 | 27 | −571.926996 | −886.866300[f] | −1040.863871[g] |
| BiphCOCH$_2$COCH3 | 1.48 | 27 | −763.536873 | −1078.508623 | |
| BnCOCH$_2$COPh | 0.21 | 27 | −763.529367 | −1078.503017 | |

[a]Adamantylacetylene. [b]IBP = isobutyrophenone. [c]Carbazole. [d]CHX = cyclohexanone. [e]Benzocyclohept-1-en-3,5-dione. [f]Benzocyclohept-1,4-dien-3-one-5-OLi$_2$E. [g]Benzocyclohept-1,3-dien-5-one-3-OLi$_3$E.
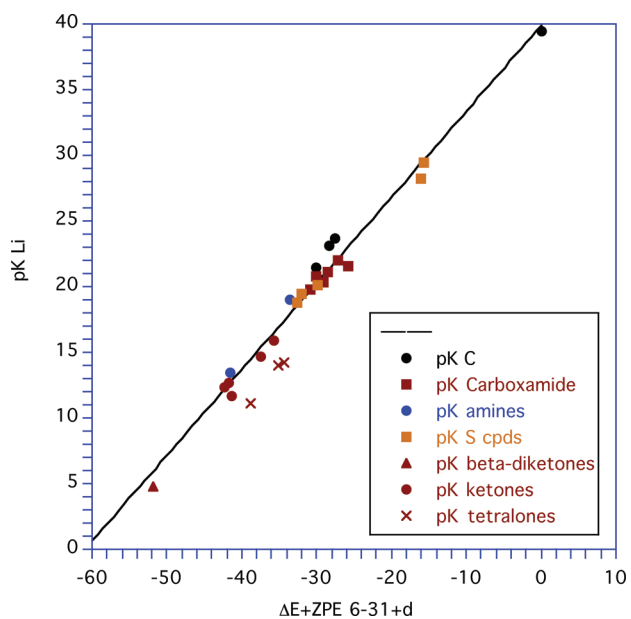


**FIGURE 3.** p$K$(Li) in THF versus $\Delta$(E + ZPE) for eq 6 and the data in Table 2. The regression line does not include the tetralone points and is given by p$K$ = 39.882 ± 0.618 + (0.654 ± 0.019)$x$; $R^2$ = 0.983. A line of the same slope through the three tetralone points has an intercept of 36.26 ± 0.12.

temperature in THF the lithium in organolithium compounds is generally tetraccordinate,[49] but complete coordination might not be so common at room temperature, the temperature at which the p$K$ determinations were made. The bond strength of the last coordinated ether is not large. For example, the computed energies of successive coordinations of dimethyl ether to phenyllithium (HF, 6-31+d + ZPE) are −20.0, −12.0, and −6.1 kcal mol$^{-1}$. When translational entropy changes are included, the addition of the last ether is much less favorable. Lande et al.[50] in a recent dynamics study of methyllithium and vinyllithium in dimethyl ether concluded that desolvation to tricoordinated species could be important. Another recent careful theoretical study of the solvation of several alkyllithiums in THF led to the conclusion that disolvation is favored at room temperature.[35] Tricoordinate lithium is not uncommon in crystal structures of lithium compounds.[51−53] With the energetics of the final

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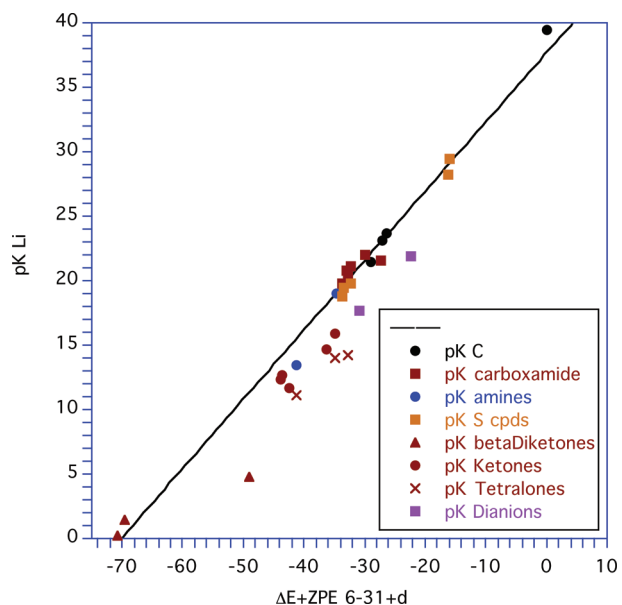(48) Wang, D. Z.-R.; Kim, Y.-J.; Streitwieser, A. J. Am. Chem. Soc. 2000, 122, 10754–10760.
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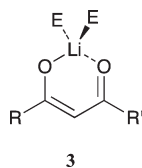
**FIGURE 4.** Plot of experimental pK(Li) for monomeric RLi versus the HF energy for eq 7. The regression line shown excludes ketones, dianions, and the W-$\beta$-diketone **1**: pK(Li) = 37.683 ± 0.493 + (0.538 ± 0.014)x; $R^2$ = 0.989. A parallel line for the ketones has an intercept of 34.10 ± 0.42.
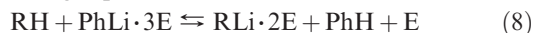
coordination being relatively small, *differences* with different substrates should be even smaller. Thus, we considered the use of only two coordinated solvents as in eq 7.

$$RH + PhLi \cdot 2E \overset{\Delta E}{\rightleftharpoons} RLi \cdot 2E + PhH \qquad (7)$$

The smaller size allows application to larger systems. Moreover, some compounds are internally coordinated and do not require three solvents. Examples are sulfones, sulfoxides, and notably, U-shaped $\beta$-diketones, which are four-coordinate with only two solvents, **3**.



**3**

If these compounds are considered in equilibrium with PhLi·3E, the resulting eq 8 no longer has a unitless equilibrium constant and the translational entropy contribution now differs substantially from zero. In addition, correction would also need to be made for the change in standard states going from the gas phase to solution.[35,54,55]

$$RH + PhLi \cdot 3E \rightleftharpoons RLi \cdot 2E + PhH + E \qquad (8)$$

A plot of pK(Li) versus $\Delta E$ (E + ZPE) for eq 7 is shown in Figure 4. The fit is very good except for ketones that now deviate almost as much as the tetralones. The two U-$\beta$-diketone lithium salts **3** (**3a**, R = Biph, R' = Me and **3b**, R = Bn, R' = Ph) fit the regression line of the other compounds, but the W-$\beta$-diketone **1** behaves more like a ketone.

The non-ketone compounds in Figure 4 have a predicted pK(Li) with an average error of ±0.6 pK units. Thermal energy corrections, as with Figure 3, have little effect on Figure 4.

Some other compounds require comment. Sulfoxides have a chiral sulfur and an attached CH$_2$ group has diastereotopic hydrogens. Such hydrogens have long been known to have different kinetic acidities[56] and to give diastereotopic lithiation and unequal pairs of trapping products at low temperature. In the computed structure of the lithium salt (**4-Li**) of benzyl phenyl sulfoxide (**4**) with three coordinating solvents (Figure 5), lithium is bonded to the sulfinyl oxygen and not at all to the $\alpha$-carbon. This carbon is virtually planar (sum of angles = 359.6°). Its two sides are still diastereotopic, but removal of either hydrogen from the parent will give the same lithium structure, and there is but one pK for the parent. In the experimental pK(Li) of the analogous phenyl biphenylylmethyl sulfoxide (**5**), both methylene hydrogens were considered as acidic and a statistical factor of 2 was used.[16] A crystal structure of the dimer of **4-Li** is available in which the lithiums are 4-coordinate by association with TMEDA and have no C−Li bonding.[57] The $\alpha$-carbons are essentially planar.

With coordination to only two ethers, however, the lithium in **4-Li** is now bound to both the sulfoxide oxygen and to carbon and there are now two structures for **4-Li·2E** that differ in energy by 1.31 Kcal mol$^{-1}$ (Figure 5). The more stable RS structure has a more pyramidal $\alpha$-carbon (sum of bond angles = 346.3° compared to SS = 351.2°) and a shorter C−Li bond (2.30 Å compared to SS = 2.37 Å). For this situation the observed acidity of **5** is associated primarily with a single hydrogen and the statistical factor of 2 is no longer appropriate; accordingly, the pK(Li) was taken as the nonstatistically corrected value of 19.8 in Figure 4. NMR[58] and spectroscopic[59] studies of PhSOCH$_2$Li in THF were interpreted in terms of an almost planar CH$_2$ group but with C−Li−O bonding essentially as in Figure 5.

The structures of lithiosulfones have also been much discussed in the literature,[60−65] but crystal structures are only available for dimers. The computed structures of the lithium salt (**5-Li**) of benzyl phenyl sulfone, **5**, are comparable to those of the sulfoxide. As summarized in Figure 6, with lithium coordinated to three ethers coordination occurs to a single sulfone oxygen and there is no C−Li bonding. The $\alpha$-carbon is somewhat nonplanar with a sum of angles of 352.9°. With two ethers, however, lithium is now also

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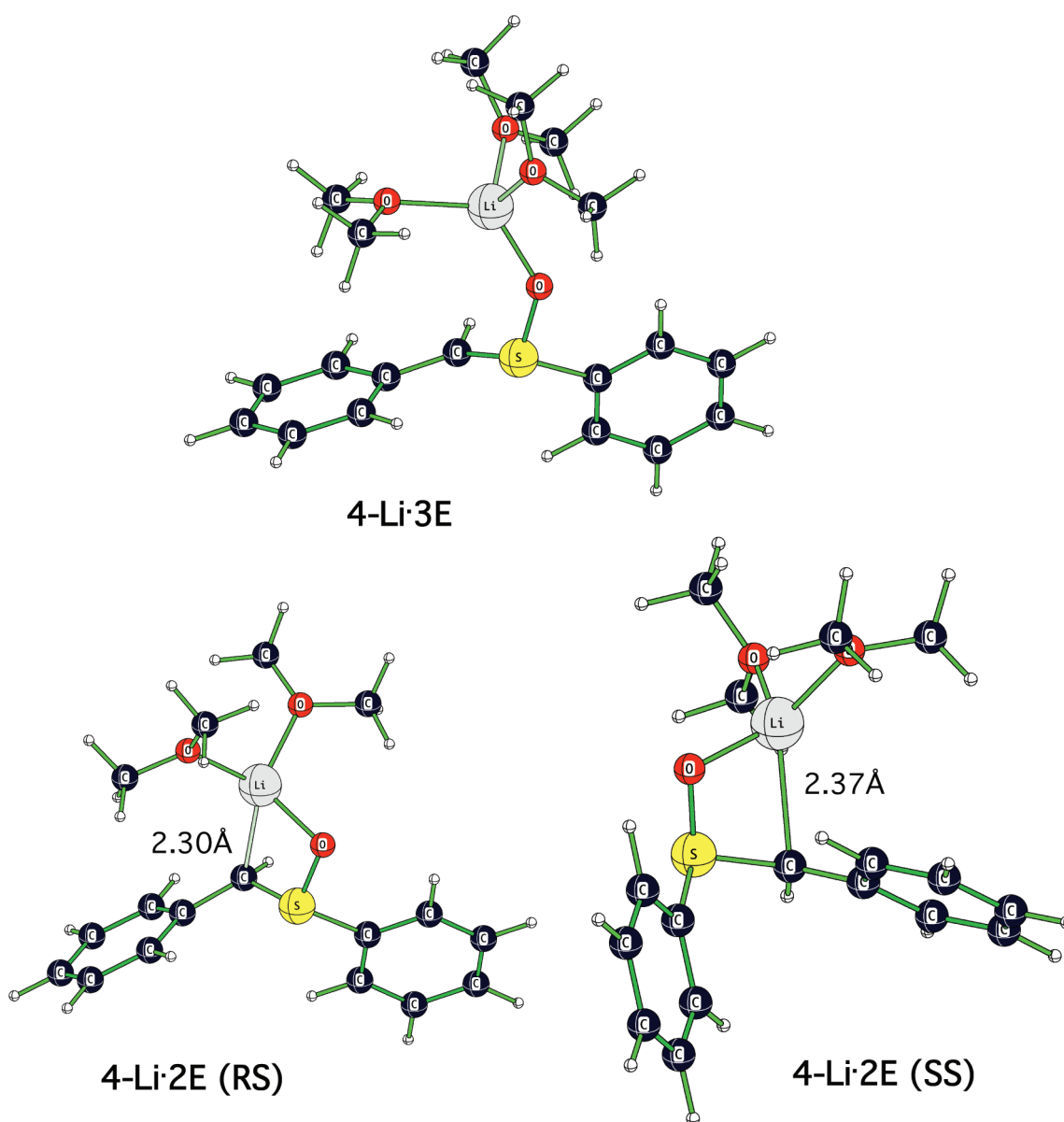(61) Bors, D. A.; Streitwieser, A., Jr. *J. Am. Chem. Soc.* **1986**, *108*, 1397–1404.
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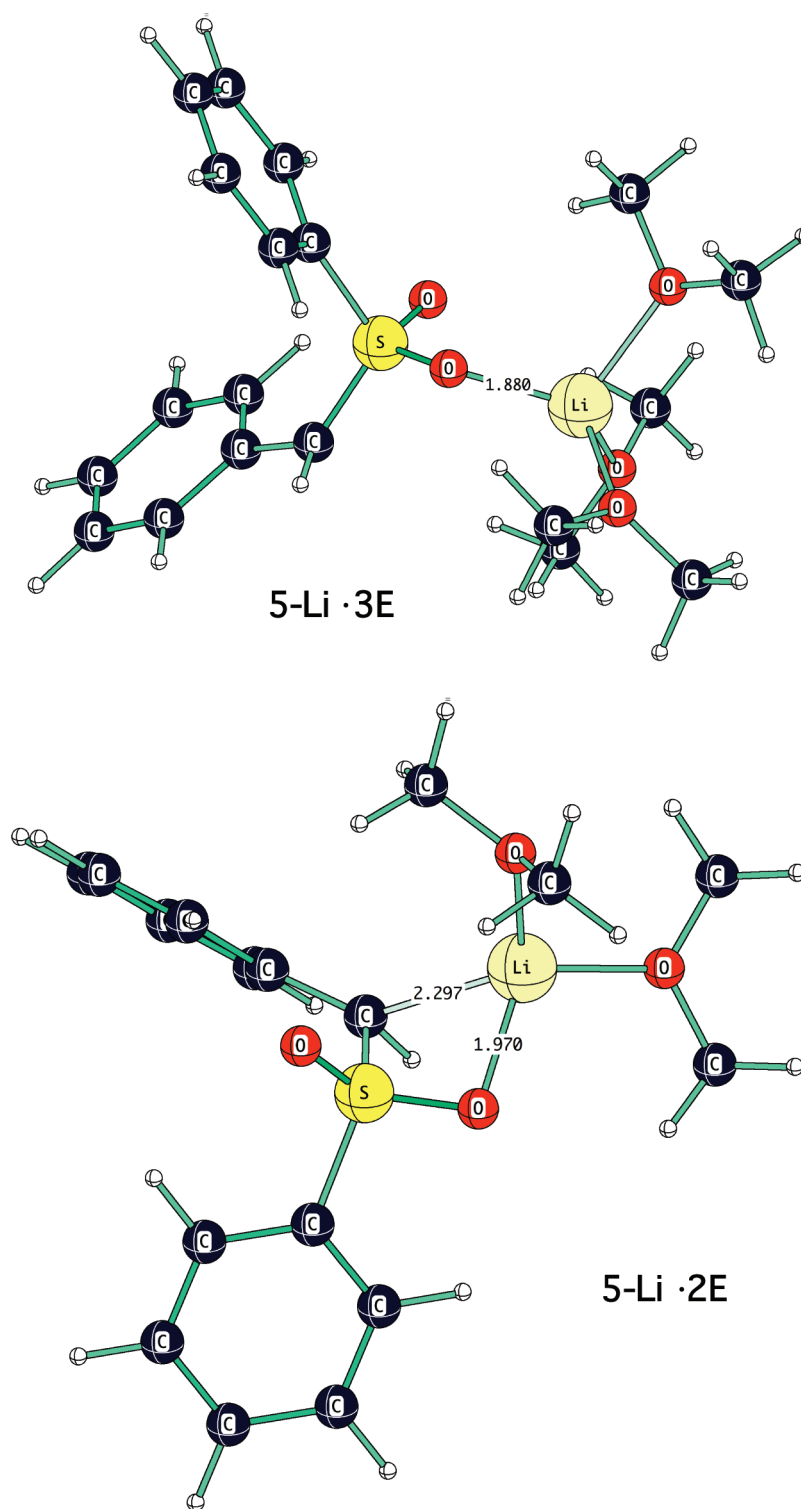
**FIGURE 5.** Structures of lithiobenzyl phenyl sulfoxide with varying degrees of coordination solvation with dimethyl ether. The RS structure is 1.31 kcal mol$^{-1}$ more stable than SS.

bonded to carbon with a normal C–Li bond length of 2.30 Å. The α-carbon is now more pyramidal with a sum of angles of 341.3°. Lithium remains coordinated to only a single sulfone oxygen.

Figure 4 also includes two new compounds, the dilithium salts of two dianions derived from deprotonation of **3a** and **3b**. Although p$K$(Li) were determined experimentally for the two monomeric dilithium compounds,[27] their structures were not established. The assumed structures based on the cyclic Li$_2$O$_2$ motif frequently found in lithium enolate dimers and suggested previously for these compounds[27] are minima on the PES (Figure 7). These two points are not far from the regression line and in fact behave more like the simple ketones. The ketone enolates follow a correlation of their own with the same slope as the normal correlation but displaced from it by −3.58 p$K$ units. With this correction, the ketones and the dilithium salts have predicted p$K$(Li)
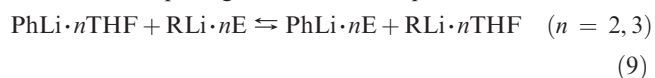
with a average error of about ±0.95 with the bicyclic ketones still contributing large deviations. The good agreement of the computed p$K$(Li)s for the dilithium salts in Figure 7 suggests that their computed structures are essentially correct.

Several further tests were made of the theory level. Most of the compounds were also computed with the 6-31++g(d,p) basis set at the 6-31+g(d) geometry, but the results are almost identical to those with the smaller basis set. MP2(FC) with the 6-31++g(d,p) basis set at the HF 6-31+g(d) geometry gives more scatter than Figure 4 and a distinctly poorer correlation (Figure S4, Supporting Information). A number of the RLi·2E systems were also computed with B3LYP (6-31++g(d,p)) at the HF 6-31+g(d) geometries, but the results show somewhat greater scatter than Figure 4. In particular, the tetralone systems are still displaced from the standard correlation (Figure S5, Supporting Information).

**FIGURE 6.** Structures of lithiobenzyl phenyl sulfone with varying degrees of coordination solvation with dimethyl ether.

Finally, we tested the use of Me$_2$O as a model for THF by running several compounds with two- and three-coordinated THF and comparing with Me$_2$O in eq 9.

$$\text{PhLi} \cdot n\text{THF} + \text{RLi} \cdot n\text{E} \rightleftharpoons \text{PhLi} \cdot n\text{E} + \text{RLi} \cdot n\text{THF} \quad (n = 2, 3)$$

$$(9)$$

If Me$_2$O were a perfect model for THF, the energy change in eq 9 would be zero. The results for six systems summarized in

Table 3 show small energy changes of generally a few tenths of a kcal mol$^{-1}$.

We conclude that within the limits of eqs 6 and 7 HF calculations at the 6-31+g(d) level provide a satisfactory correlation of CIP organolithium equilibria. The perhaps surprising superiority of HF methods might well be associated with interactions with lithium and with the coordinating solvents being largely electrostatic interactions
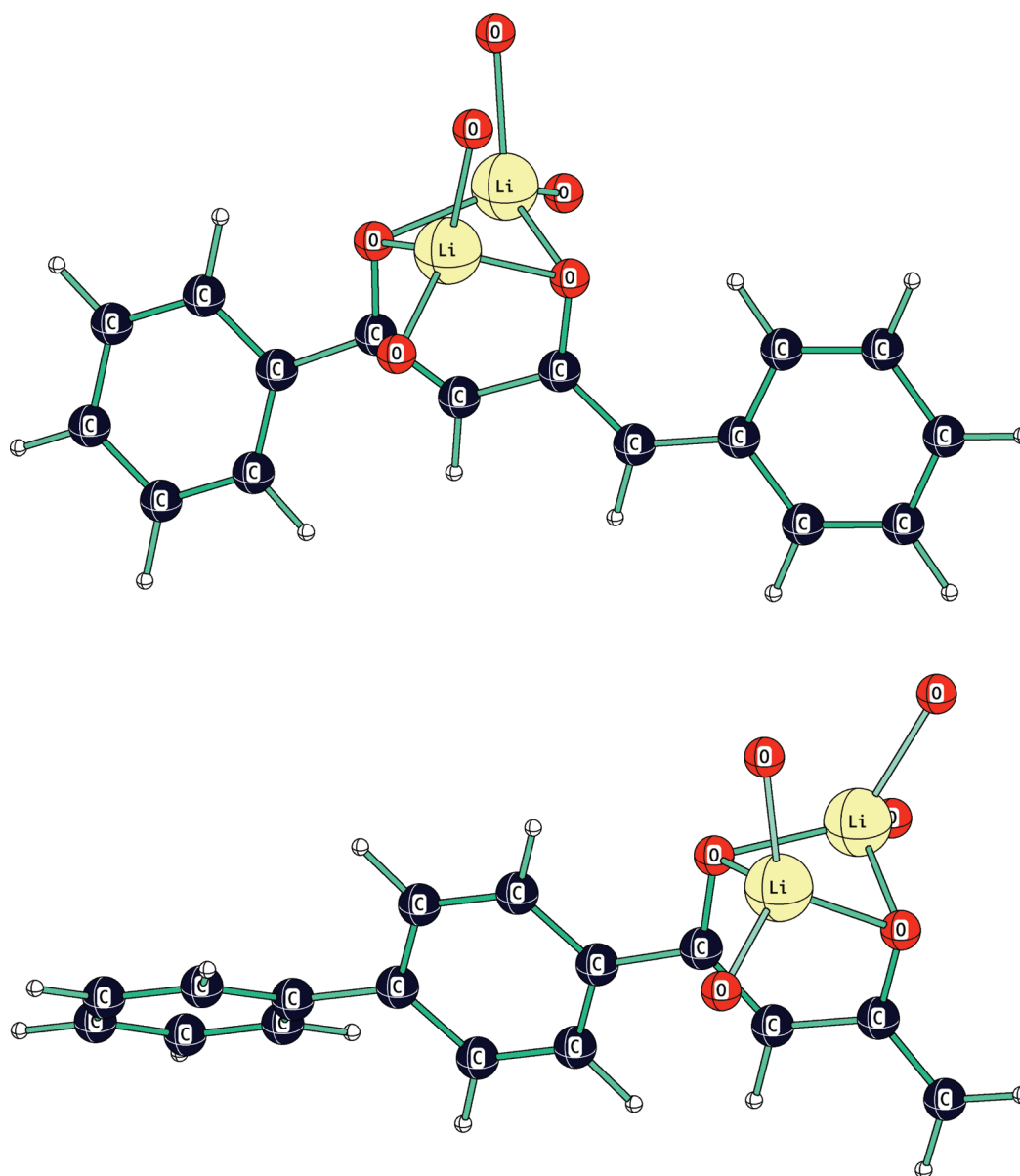
**FIGURE 7.** Optimized structures at HF 6-31+g(d) of the dilithium salts derived from **3**. Each lithium is coordinated to two dimethyl ethers; the methyl groups are not shown for clarity.

**TABLE 3.** Energy Changes in kcal mol$^{-1}$ for Equation 9

| R | $n = 2$ | $n = 3$ |
|---|---|---|
| Ph$_2$CHCONMe$_2$ | −0.075 | 0.416 |
| carbazole | −0.564 | −0.539 |
| Phcyclohexanone | 0.240 | 0.024 |
| tetralone | −0.348 | 0.729 |
| acetaldehyde | −0.018 | 0.063 |
| C$_6$H$_2$F$_4$ | −0.636 | −0.868 |

**TABLE 4.** Predicted p$K$'s Using the Correlation in Figure 3

| compound | calculated p$K$ (per H) |
|---|---|
| isobutyrophenone | 16.7 |
| phenyl-($N$,$N$-dimethylacetamide) | 20.2 |
| tetralone | 14.4 |
| 2-phenyltetralone | 11.4 |
| 2-benzyltetralone | 13.8 |

of closed shells with only small differences in covalent character. The best correlations were found using E + ZPE; that is, the energies at 0 K. Correction to room temperature using the computed thermal energies have virtually no net effect; the corrections are much the same for both sides of eqs 6 and 7. Use of the computed $\Delta G°$ was not successful, probably because of the important role of

low frequency vibrations and their anharmonicity on entropy, as discussed above.

The correlations found can obviously be used to calculate the p$K$(Li)s of other compounds, and in subsequent papers we will apply these correlations to the experimental p$K$'s determined by others in which the possible role of aggregation was not considered. We will also show how the computed p$K$'s can be combined with experimental kinetic acidities to obtain mixed

experimental–theoretical Brønsted correlations.[66] For the present, we will limit ourselves here to the prediction of several compounds related to those in Table 2 but lacking a phenyl group required for the p$K$ measurements. These results are summarized in Table 4. The tetralone compounds made use of the special tetralone correlation in Figure 3. As expected, a phenyl substituent has an effect of only a few tenths p$K$ unit. Note that these p$K$'s refer to the monomer; these organolithium compounds are expected to be highly aggregated in THF solution.

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**Supporting Information Available:** Tables of computational results and additional figures. This material is available free of charge via the Internet at http://pubs.acs.org.