through both pieces. The apparatus was tipped so that the Me_2SO covered the hydride and the glass frit. Evolution of hydrogen began immediately. After the reaction was over, the solution was filtered through the frit directly into a storage flask fitted with a three way stopcock by evacuating the bottom of the fritted funnel. The storage flask was removed from the apparatus, again with a flow of argon going through both pieces. The flask was stoppered, degassed three times, and wrapped in aluminum foil. Solutions (~100 mM) prepared in this manner were colorless.

Indicator Method. The details of the method of measurement using indicator anions have been previously described.[5] Very briefly, it involves three parts: (1) determination of the extinction coefficient of the indicator anion via a Beer's law plot through the addition of a few aliquots of an Me_2SO solution of the indicator acid (pK_a known) to a solution of potassium dimsyl ($[K^+CH_2SOCH_3^-]$ = 1-5 mM) in Me_2SO; (2) addition of an excess of indicator acid and measurement of the absorbance of the solution to determine the anion concentration; (3) addition of aliquots of an Me_2SO solution of the acid of unknown pK_a and measurement of the absorbance. Modifications to this method that have been developed in the past several years include a change in the method of preparation of potassium dimsyl (see above) and the use of extra precautions for compounds in the high pK_a region (see below). As mentioned above, argon is now used in place of nitrogen, a modification which seems to decrease the amount of oxygen coming into contact with the oxygen-sensitive carbanion solutions during runs. (The presence of oxygen during titrations makes itself evident by absorbance changes with time or after shaking the spectrophotometric cell.) The indicator anions in this study are particularly sensitive to oxygen, so only freshly distilled Me_2SO and freshly prepared $CH_3SOCH_2^-K^+$ were used. Normal acid-base equilibria in Me_2SO (eq 1) are insensitive to temperature but secondary equilibria (e.g., eq 2 and 3) can be, so the cell compartment in the Cary 14 spectrophotometer was thermostatted to 25 ± 1 °C with a Beckman recirculating water bath.

Enough of a solution of cryptand in Me_2SO was added to the spectrophotometric cell so that it was in excess of the amount of potassium for all of the alcohol and water runs.

Acknowledgment. We are grateful to the National Science Foundation and to the donors of the Petroleum Research Fund, administered by the American Chemical Society, for support of this work.

Registry No. Me_2SO, 67-68-5; H_2O, 7732-18-5; MeOH, 67-56-1; EtOH, 64-17-5; *i*-PrOH, 67-63-0; *t*-BuOH, 75-65-0.

Ion-Pair Association Constants in Dimethyl Sulfoxide

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A method of detecting ion-pair formation between the anions of organic acids and metal cations in conjunction with acidity measurements is presented. The equilibrium between an indicator anion which does not ion pair, an anion which does ion pair, and their conjugate acids is perturbed by the addition of a metal iodide salt. The association constant for ion-pair formation is derived from the spectrophotometric data. This allows pK_a values to be corrected for ion pairing. Most organic anions do not form ion pairs with potassium cations in dilute (1-5 mM) dimethyl sulfoxide solution. For β-diketones, conformational changes were found to produce large effects, not only on the association constants, K_{as}, for their anions but also on their acidities in Me_2SO relative to H_2O. With K^+ ion, β-keto enolate ions fixed in the *E,E* (*W*) conformation chelated least, acyclic β-keto enolate ions (*E,Z*) chelated more, and β-keto enolate ions fixed in the *Z,Z* (*U*) conformation chelated most. With these anions the size of K_{as} varied with the nature of the cation in the order $K^+ < Na^+ < Li^+$. For $PhCO_2^-$ and $H_2C{=}NO_2^-$ anions the differences in K_{as} were smaller, and the order was $K^+ < Li^+ < Na^+$. The chelating ability of $[\mathrm{O}\overset{..}{-}\mathrm{C(R)}\overset{..}{-}\mathrm{X}]^-$ anions toward K^+ ion was found to be $RCO_2^- > RCONR^- > RCOCR_2^-$. The anion derived from a cyclic carboxamide, 1-azacyclohexan-2-one, chelated with K^+ ion much more strongly than did an acyclic analogue anion, $EtCONEt^-$. MeO^- ion pairs strongly with K^+ whereas the equally basic carbanions, $PhC{\equiv}C^-$ and thiazolide ion, do not.

The formation of ion pairs in organic chemistry is widespread. Several recent reviews have been devoted to their detection and to studies of their spectroscopic properties, chemical reactivity, and association constants.[1-3] The existence of two types of ion pairs,[1] contact (or tight) and solvent separated (or loose), has been demonstrated by a variety of kinetic and spectroscopic techniques. Many of the factors which influence the formation of these two types of ion pairs have been delineated. These include solvent, temperature, anion and cation identity and structure, and the addition of ion-complexing agents. Most of the ions which have been studied are hydrocarbon anions and radical anions. There have been few studies of anions with functional groups such as carbonyl, nitrile, or nitro.

Parallel to the development of these physical chemical studies, many organic chemists have been varying the conditions of synthetic reactions to promote or suppress ion-pair formation in order to improve yields, speed up reactions, or favor product ratios or stereochemistry.[4] The primary result of changing from a weakly dipolar nonhydroxylic solvent such as tetrahydrofuran to a strongly dipolar nonhydroxylic solvent such as hexamethyl-

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phosphoramide or dimethyl sulfoxide is a decrease in the amount or strength of ion–counterion interaction. Variation of counterion, addition of a salt, and addition of a cation complexing agent such as crown ether are other techniques that have been used extensively in synthetic organic chemistry.

With one notable exception there has been little overlap between the qualitative observations of synthetic chemists and quantitative and structural studies of physical and physical organic chemists. The exception is the alkylation of ambident nucleophiles, especially β-keto enolate anions.[5-7] It has been shown that one of the factors which favor O-alkylation is a decrease in the interaction between the enolate and the counterion.[5] Several workers have correlated product ratios with the amount of ion-pair formation, the latter having been measured by conductivity techniques.[6] In the most complete work to date in this area, Arnett and co-workers have used conductance, IR, UV, and ^1H and ^{13}C NMR spectroscopy, calorimetry, and kinetics to study the structure and alkylation of several β-keto enolate anions.[7]

β-Keto enolate ions have received extensive study because of their ease of generation in aqueous solution and their strong chelating properties. On the other hand, relatively little is known about the chelating ability of anions that are too strongly basic to be generated in appreciable quantity in hydroxylic solvents. For example, little is known about the association constants of enolate ions with cations because ketones are too weakly acidic to allow enolate ions to be generated in high concentrations in the hydroxylic solvents generally used to measure such association constants.[8] This is true also for the conjugate bases of most other carbon acids. The development of an acidity scale in the strongly dipolar nonhydroxylic ("aprotic") solvent dimethyl sulfoxide that allows absolute equilibrium acidity measurements to be made for acids as much as 20 pK_a units weaker than those measured in water has provided a method of generating known concentrations of strongly basic anions.[9] There is evidence to indicate that, in general, there is little or no ion pairing between most of these anions with potassium ion in the dilute solutions used for these acidity measurements. β-Keto enolate ions have been shown to chelate with K^+ ion under these conditions, however, and ion association constants with K^+, Na^+, and Li^+ in Me_2SO have been measured by conductance methods.[7] It seemed likely that our indicator pK_a method could be modified so as to provide a simple test for ion pairing and a method for measuring association constants, K_{as}. This has been accomplished, and we now present the results and discuss the factors important in ion-pair association, namely, anion structure, conformation, and charge delocalization.

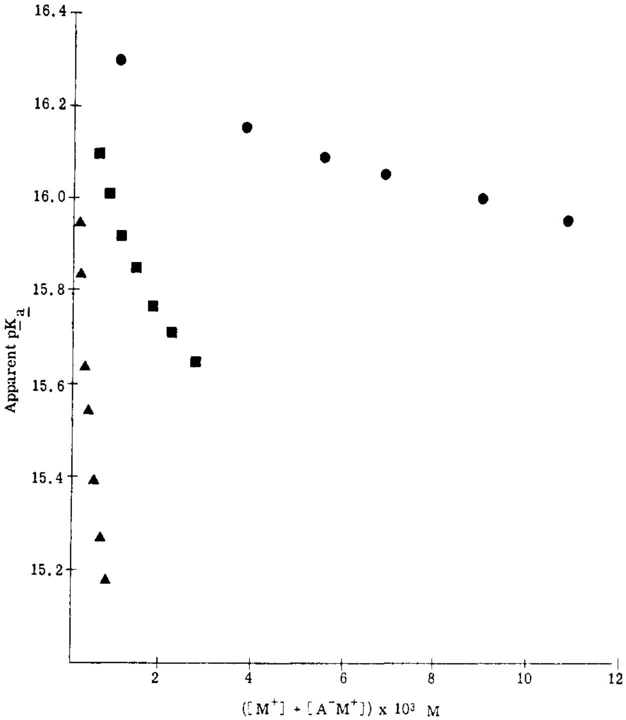

Figure 1. Plots of apparent pK_a (eq 4) vs. total cation concentration for diethyl malonate: ●, $M^+ = K^+$, calculated pK_a = 16.38 ± 0.00, log K_{as} = 2.32 ± 0.02, R^2 = 0.999; ■, $M^+ = Na^+$, calculated pK_a = 16.38 ± 0.04, log K_{as} = 3.34 ± 0.02, R^2 = 0.993; ▲, $M^+ = Li^+$, calculated pK_a = 16.39 ± 0.00, log K_{as} = 4.57 ± 0.01, R^2 = 0.997. Data are from Table I.

Method and Results

Indicator methods have been particularly useful for the study of organic acids and their anions in nonaqueous solvents.[9-12] Usually the purpose of these methods has been to find the acidity of an acid by measuring the equilibrium constant between the acid and an indicator acid of known acidity. In nonpolar or weakly dipolar solvents such as benzene, cyclohexylamine, or tetrahydrofuran an ion-pair equilibrium (eq 1) is established,

$$M^+In^- + H\text{–}A \rightleftarrows HIn + M^+A^- \tag{1}$$

$$In^- + H\text{–}A \rightleftarrows HIn + A^- \tag{2}$$

$$M^+ + A^- \rightleftarrows M^+A^- \tag{3}$$

while in strongly dipolar solvents such as dimethyl sulfoxide an equilibrium free of cation effects (eq 2) is usually established. For those anions that do pair with cations in Me_2SO an additional equilibrium (eq 3) is also present. Ordinarily eq 3 can be ignored because of the strong solvating power of Me_2SO for K^+ ion and the dilute solutions used. If ion association occurs, however, eq 3 becomes important, and eq 2 is perturbed. The apparent (observed) pK_a will then be lower than the true value.

Ion-pair formation can be detected by adding five to ten increments of a standard solution of K^+I^- in Me_2SO to the equilibrium established in the pK_a measurement. The concentration of indicator anion, In^-, is measured spectrophotometrically after each addition. A drop in absorbance greater than that which can be due to simple dilution immediately shows that eq 3 is significant. This constitutes a qualitative test for ion pairing, which is now used routinely in our laboratory. It can be extended to test also for ion pairing involving Na^+, Li^+, or other cations.

An apparent equilibrium constant defined by eq 4, can be measured as a function of total cation concentration by using the data obtained in the ion-pairing test just de-

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$$\text{apparent } K_a = \frac{[\mathrm{HIn}]([\mathrm{A}^-] + [\mathrm{A}^-\mathrm{M}^+])}{[\mathrm{In}^-][\mathrm{HA}]} \qquad (4)$$

scribed. Plots of $[M^+] + [A^-M^+]$ vs. the apparent pK_a for K^+, Na^+, and Li^+ and the anion derived from diethyl malonate are shown in Figure 1.

A general least-squares procedure was used to calculate the true equilibrium constants for eq 2 and 3. We assume that neither the indicator anion, which is a delocalized hydrocarbon anion (usually a substituted fluorenyl anion) nor the iodide anion is ion paired with the cation. At the low ion concentrations we are using (10^{-5} to 2×10^{-2} M) this is a safe assumption.[13,14] Activity coefficients for neutral species were assumed to be unity, and for ions they were calculated by using simple Debye–Hückel theory. A value of 4 Å was chosen for the adjustable parameter a. The calculated association constants are insensitive to the value of a. Even use of the limiting law ($a = 0$) causes very little change in the calculated values. As an example, for the potassium ion run in Figure 1, changing a from 4 to 0 Å changes the calculated pK_a from 16.37 to 16.38 and $\log K_{as}$ from 2.32 to 2.35. The data for Figure 1 are shown in Table I. Also shown are the apparent pK_a's calculated from the least-squares pK_a and K_{as}. The agreement is excellent.

Several internal checks were performed to make sure that the method was working properly. In one run with acetylacetone and sodium ion, the calculated pK_a was 13.35, and the apparent pK_a was 12.85 at a Na^+ concentration of 6.6 mM. Addition of 1 equiv of [2.2.2]cryptand[15] resulted in an apparent pK_a of 13.35, indicating that the cryptand effectively eliminated ion pairing of Na^+ with $(CH_3COCHCOCH_3)^-$. For all of the lithium runs the anions were generated from $K^+CH_3SOCH_2^-$ in a concentration range where there was very little ion pairing of K^+ ion with the conjugate base. Thus a true pK_a could be obtained before the lithium titration. In all cases this pK_a value was within ±0.03 of the pK_a calculated from the lithium titration.

Two of the acids in Table II, dimedone and benzoic acid, required a special technique because of the presence of equilibrium 5 in addition to equilibria 2 and 3. This type

$$A^- + HA \rightleftarrows A^- \cdots HA \qquad (5)$$

of equilibrium is common in nonhydroxylic solvents for oxygen acids (dimedone is >90% enol in Me_2SO).[16] To study these acids, we calculated equilibrium constants for eq 2 and 5 from an acid titration of a solution of indicator and indicator anion in the presence of 1 equiv of [2.2.2]-cryptand to complex the potassium counterion. With these two equilibrium constants known, an association constant could be calculated for each subsequent addition of alkali metal iodide.

The pK_a values and association constants for these two acids and others that do not hydrogen bond to their conjugate bases are shown in Table II. For dimedone the calculated value of the homohydrogen bonding constant, $\log K_{AHA}$ (eq 5), is 2.20 ± 0.20, and for benzoic acid it is 1.90 ± 0.09. The latter agrees with the previously reported value of 1.8 in Me_2SO.[16]

The β-dicarbonyl compounds in Table II can exist as either the keto or enol tautomers. In Me_2SO solution, acetylacetone is known to be 40% keto.[17] The percent keto configuration for the other compounds was determined from their ^1H NMR spectra in CD_3SOCD_3. The results are dibenzoylmethane (<10), diethyl malonate (>90), 3-methyl-1,8-decalindione (<10), and dimedone (<10).

Discussion

Acidities. The acidities in Me_2SO, corrected for ion pairing, are shown in Table II. The corrections depend on the concentration of base used in the pK_a measurement, as well as on K_{as}. For example, with 2.90 mM base in the cuvette and K^+ cation the correction for the pK_a of CH_3NO_2 is 0.05 pK_a unit. With Na^+ ion, where K_{as} is almost an order of magnitude larger, the correction for 2.94 mM base is 0.26 pK_a unit but drops to 0.08 pK_a unit with 0.42 mM base. The corrected pK_a's for CH_3NO_2 and $(CH_3CO)_2CH_2$ are within experimental error of those reported in earlier publications, but that for $(PhCO)_2CH_2$ is about 0.3 pK_a unit higher than that reported earlier. The correction for the strongest chelating β-keto enolate ion (**5**) amounts to about 0.6 pK_a unit.

Table III shows a comparison of these acidities with those in water and with the heats of deprotonation by the lyate ion $(CH_3SOCH_2^-)$. The greater acidity of these β-diketones in water than in Me_2SO is probably due primarily to the stabilizing influence of hydrogen bonding on the anion in water. For a monoketone, where the charge is largely localized on oxygen, ΔpK_a is probably about 8.[18] For the acyclic β-diketones, we see from Table III that ΔpK_a averages about 4. The lesser stabilization of these anions in water is due to the greater delocalization of charge over the $[O\cdots C\cdots C\cdots C\cdots O]^-$ system.

The importance of conformation in prescribing acidities in Me_2SO and water and in prescribing the difference in acidities in these two solvents can be illustrated by comparing the acidities of the conjugate acids of **5** and **6** [4-methyl-1,8-decalindione and 5,5-dimethylcyclohexane-1,3-dione (dimedone), henceforth designated as **5-H** and **6-H**, respectively] with those of their acyclic analogues, 3-methylpentane-2,4-dione (**2-H**) and pentane-2,4-dione (**1-H**). Both **5-H** and **6-H**, as well as their acyclic analogues, **2-H** and **1-H**, exist primarily as enols in Me_2SO solution. Enol **5-H** is less acidic than its acyclic analogue, enol **2-H**, in Me_2SO by 3.1 pK_a units, whereas enol **6-H** is *more* acidic in Me_2SO than its analogue, enol **1-H**, by 2.1 pK_a units. One factor dictating the order of these relative acidities is the fixed (*U*) conformation of enol **5-H** which enhances intramolecular hydrogen bonding (acid weakening) and the fixed (*W*) conformation of enol **6-H** which prevents intramolecular hydrogen bonding (acid strengthening). A second factor is the fixed orientation of the C–O dipoles, which is repulsive (between the oxygen ends of the dipoles) in anion **5** (acid weakening) and repulsive in enol **6-H** (between the carbon ends of the dipoles) (acid strengthening).

The acidities of **5-H** and **6-H** are greater in water than in Me_2SO by 6.3 and 5.4 pK_a units, respectively (ΔpK_a in Table III). Nitromethane and benzoic acid are more acidic in water than in Me_2SO by similar amounts (ΔpK_a's = 6.6 and 6.3). As mentioned above, these effects are due primarily to stabilization of anions **5**–**8** by hydrogen bonding in water. In anions **1**–**3** these hydrogen bonding effects

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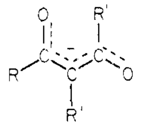
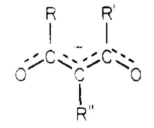
Table I. **Experimental Data and Calculated Apparent pK_a's for Three Runs of Diethyl Malonate**

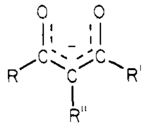
cation	point no.	$10^3([A^-] + [A^-M^+])$, M	$10^3([M^+] + [A^-M^+])$, M	app pK_a (exptl)^a	app pK_a (calcd)^{a,b}	% ion pairing^c
K^+	1	0.229	1.036	16.298	16.300	16.1
	2	0.274	3.817	16.160	16.160	39.0
	3	0.293	5.473	16.098	16.100	47.1
	4	0.304	6.873	16.060	16.056	51.5
	5	0.322	9.029	16.000	15.999	57.8
	6	0.334	10.946	15.953	15.956	62.2
	7	0.344	13.198	15.909	15.912	65.8
Na^+	1	0.232	0.537	16.103	16.107	46.9
	2	0.252	0.832	16.011	16.004	57.0
	3	0.271	1.115	15.922	15.924	65.0
	4	0.287	1.407	15.848	15.854	70.5
	5	0.304	1.839	15.761	15.768	75.9
	6	0.313	2.222	15.707	15.705	78.7
	7	0.324	2.709	15.642	15.636	81.8
Li^+	1	0.114	0.124	15.940	15.944	64.3
	2	0.132	0.173	15.829	15.825	72.3
	3	0.166	0.278	15.629	15.625	82.6
	4	0.183	0.336	15.534	15.534	86.0
	5	0.210	0.457	15.379	15.382	90.2
	6	0.230	0.580	15.252	15.259	92.7
	7	0.242	0.703	15.160	15.158	94.1

^{a} Apparent pK_a's are defined in eq 3. Apparent pK_a = pK_{HIn} + log K, K as defined in eq 3. ^{b} Calculated from the least-squares pK_a and K_{as} given in Figure 1. ^{c} $([A^-M^+]/([A^-] + [A^-M^+]))100$.

may be smaller because these anions exist in conformations differing from the fixed *Z,Z* (*U*) and *E,E* (*W*) conformations of **5** and **6**. Anions **1**–**3** have three possible conformations (**9**–**11**). When R = R′ = Me and R″ = H, the *E,Z* (*S*) conformation (**9**) has been shown by NMR to be favored in methanol.[20] In this conformation both the repulsion between the two partially charged oxygen atoms and the steric interactions between R and R′ are minimized. Anions **2** and **3** are also likely to prefer this conformation. The smaller ΔpK_a values found for the conjugate acids of **1**–**3** can be attributed in part to conformational effects in the anions leading to less hydrogen bonding and in part to differences in other structural effects such as keto–enol ratios in the two media.


9, *E,Z* (*S*)


10, *E,E* (*W*)

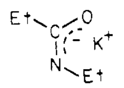

11, *Z,Z* (*U*)

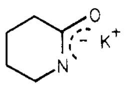
Substitution of an α-Me group into **1** [MeCOCH$_2$COMe → MeCOCH(Me)COMe] causes a 1.8 pK_a unit decrease in acidity. Acid-weakening (and acid-strengthening) Me effects may have a number of different causes and are therefore difficult to interpret.[21] We note, for example, that similar acid-weakening Me effects are observed in MeCH(CN)$_2$ and MeCH(SO$_2$R)$_2$ substrates,[21] which differ rather markedly from **2** in structure.

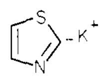
Ion Pairing. The various conformations of β-keto enolate anions have important influences on the amount of ion pairing, as well as on the acidities of their conjugate acids. The coupling of acidity and association-constant measurements can reveal important information about the charge distributions in these anions.

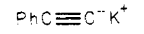
The ion association constants measured by the indicator method agree with those determined by the conductance method to better than 0.2 logarithmic unit for K$^+$ and Na$^+$ complexing with [PhCOCHCOPh]$^-$ and [MeCOC(Me)COMe]$^-$ (Table II). Neither our method nor the conductance method distinguishes between solvent-separated ion pairs and contact ion pairs. It seems likely that we are observing the effects of contact ion pairs since at these low concentrations ion association in Me$_2$SO with K$^+$ occurs only with anions that interact in a specific way, particularly by chelation. Spectroscopic evidence on metal β-keto enolates supports this point of view.[7a]

Structural Effects on Ion Pairing. The data in Table II show that ion association is significant in dilute Me$_2$SO solution for β-keto enolate ions and for other anions wherein chelates can be formed involving K$^+$, Na$^+$, or Li$^+$ and two oxygen atoms bearing partial negative charges that have the proper geometry. Anions wherein the negative charge is localized on oxygen are also known to associate with alkali metal cations. For *t*-BuO$^-$ the K_{as} values in Me$_2$SO are 270, 10^6, and 10^8 for K$^+$, Na$^+$, and Li$^+$, respectively; MeO$^-$ and other alkoxide ions associate more strongly.[23] Ion pairing between K$^+$ and the enolate ions derived from acetophenone, cyclohexanone, or phenol were not sufficient to be revealed by our test, however (K_{as} < 20). The anion derived from an open-chain carboxamide, *N*-ethylpropionamide (**12**), was small ($K_{as} \simeq 20$), but that for the anion derived from a six-membered cyclic carboxamide analogue (**13**), where a cis structure is fixed by the ring, was appreciable (K_{as} = 350).[25] This large difference is understandable since acyclic carboxamides are known to exist in a conformation wherein the R groups are trans to one another.[26]


12 ($K_{as} \simeq 20$)


13 (K_{as} = 350)


14 (K_{as} < 20)


15 (K_{as} < 20)

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(20) (a) Raban, M.; Noe, E. A.; Yamamoto, G. *J. Am. Chem. Soc.* **1977**, *99*, 6527. (b) Raban, M.; Yamamoto, G. *Inorg. Nucl. Chem. Lett.* **1976**, *12*, 949.

(21) Bordwell, F. G.; Bartmess, J. E.; Hautala, J. A. *J. Org. Chem.* **1978**, *43*, 3095.

(22) Hogen-Esch, T. E.; Smid, J. *J. Am. Chem. Soc.* **1966**, *88*, 318.

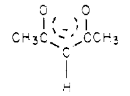
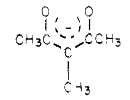
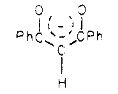
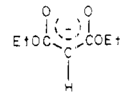
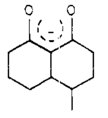
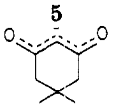
(23) Exner, J. H.; Steiner, E. C. *J. Am. Chem. Soc.* **1974**, *96*, 1782.

(24) Cox, B. G.; Hedwig, G. R.; Parker, A. J.; Watts, D. W. *Aust. J. Chem.* **1974**, *27*, 477.

(25) Algrim, D., unpublished results.

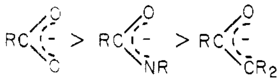
(26) Robin, M. B.; Bovey, F. A.; Basch, H. "The Chemsitry of Amides"; Zabicky, J., Ed.; Wiley Interscience: New York, 1970; Chapter 1, p 19.

Table II. Ion-Pair Association Constants for Various Anions with Alkali Metal Cations in Dimethyl Sulfoxide Solution at 25 °C

anion	cation (M^+)	runs[a]	pK_a of conj acid[b]	log K_{as}
 **1**	K^+	4	13.33 ± 0.05	1.39 ± 0.09
	Na^+	4		2.60 ± 0.06
	Li^+	3		4.77 ± 0.01
 **2**	K^+	3	15.07 ± 0.06	0.79 ± 0.10 (0.89 ± 0.22)[c]
	Na^+	3		1.52 ± 0.09 (1.58 ± 0.19)[c]
	Li^+	3		3.38 ± 0.14 (2.76 ± 0.08)[c]
 **3**	K^+	6	13.36 ± 0.06[e]	2.76 ± 0.07 (2.72 ± 0.04)[d]
	Na^+	3		3.97 ± 0.04 (3.79 ± 0.10)[d]
	Li^+	4		5.86 ± 0.11
 **4**	K^+	4	16.37 ± 0.06	2.31 ± 0.02
	Na^+	4		3.30 ± 0.06
	Li^+	4		4.56 ± 0.07
 **5**	K^+	5	18.16 ± 0.07	3.39 ± 0.04
	Na^+	5		4.93 ± 0.14
	Li^+	3		>7[f]
 **6**	K^+	3	11.16 ± 0.02	<1
	Na^+	3		1.05 ± 0.14
	Li^+	3		1.58 ± 0.03
$CH_2{=}NO_2^-$ **7**	K^+	5	17.23 ± 0.04	1.74 ± 0.03
	Na^+	6		2.70 ± 0.03
	Li^+	4		2.25 ± 0.02
$PhCO_2^-$ **8**	K^+	2	11.02 ± 0.03	1.68 ± 0.02
	Na^+	3		2.31 ± 0.01
	Li^+	3		2.21 ± 0.03

[a] Each run was a 5-10 point titration using M^+I^- as the cation source. Except for dibenzoyl methide, the acids were equilibrated with two indicator acids. [b] The pK's were calculated from the known pK's of the indicator acids and the measured equilibrium constants for eq 2. [c] Arnett, E. M., private communication. [d] Reference 7a. The lithium salt of this anion was reported as having negligible conductance. [e] The previously reported value of 13.1 [Bordwell, F. G.; Van Der Puy, M.; Vanier, N. R. *J. Org. Chem.* 1976, *41*, 1883] is low by 0.26 pK_a unit due to failure to correct for ion pairing. [f] The association constant is so high that even in the 10^{-5} M range there are an insignificant number of free ions. Evidence for triple-ion formation (LiA_2^-) was obtained at higher concentrations by adding less than 1 equiv of LiI and observing a decrease in absorbance greater than could be due to quantitative formation of LiA.

For $[O\cdots C(R)\cdots X]^-$ anions the order of chelating ability is



This suggests that the order of chelating ability of these ambident ions with K^+ is determined by the electronegativity of the atom conjugated with oxygen, i.e., C < N < O. Judging from the K_{as} constants determined for $PhCO_2^-$ and $H_2C{=}NO_2^-$ (Table II), we can expect the order of chelating ability of these anions with alkali metal cations to be $K^+ < Li^+ < Na^+$. (These cation effects are discussed further in a latter section.)

Table III. Acidity Data for the Conjugate Acids of Anions 1-8

anion	pK_a (Me_2SO)	ΔH_D (Me_2SO),[a] kcal/mol	pK_a (H_2O)	ΔpK_a[k]
1	13.3	−31.7[b]	8.9[e]	3.8
2	15.1	−30.3[c]	10.9[f]	3.6
3	13.4	−33.2[c]	9.8[g]	3.0
4	16.4		15.2[h]	0.6
5	18.2		11.3[i]	6.3
6	11.2	−33.9[b]	5.2[e]	5.4
7	17.2		10.2[j]	6.6
8	11.0	−31.5[d]	4.2[j]	6.2

[a] Enthalpy of deprotonation by the lyate anion. [b] Arnett, E. M.; Johnston, D. E.; Small, L. E. *J. Am. Chem. Soc.* 1975, *97*, 5598. [c] Reference 7h. [d] Arnett, E. M.; Small, L. E.; Oancea, D.; Johnston, D. *J. Am. Chem. Soc.* 1976, *98*, 7346. [e] Schwarzenbach, G.; Lutz, K. *Helv. Chim. Acta* 1940, *23*, 1162. [f] Laloi, L.; Rumpf, P. *Bull. Chem. Soc. Fr.* 1961, 1645. [g] Eistert, B.; Markel, E.; Reiss, W. *Chem. Ber.* 1954, *87*, 1513. [h] Vermesse-Jacquinot, C.; Schaal, R.; Rumpf, P. *Bull. Chem. Soc. Fr.* 1960, 2030. [i] Reference 34. [j] Bell, R. P. "The Proton in Chemistry"; Cornell University Press: Ithaca, NY, 1959. [k] Corrected to an absolute basis by using the equation $\Delta pK_a = pK_a(Me_2SO) - pK_a(H_2O) - 0.6$ (see following paper for a discussion).

The failure of the localized carbanions **14** and **15** to ion pair strongly with K^+ demonstrates the lesser ion-pairing ability of a carbanion as compared to an oxide ion, since **14** and **15** have basicities comparable to MeO^-, yet only the latter ion pairs strongly.[23] (The pK_a values in Me_2SO for the conjugate acids of **14**, **15**, and MeO^- are 29.4,[27] 28.7,[28] and 29.0,[29] respectively.)

The presence of the strongly basic carbanions **14** and **15** in dissociated form attests to the powerful solvating ability of Me_2SO for cations. Judging from single-ion free energies of transfer from water,[24] Me_2SO stabilizes K^+ by 2.9 kcal/mol relative to H_2O and by 4.3, 4.8, and 5.3 kcal/mol relative to propylene carbonate, acetonitrile, and MeOH, respectively. Other factors also play an important role, however, as is made clear by the fact that ion pairing is about an order of magnitude greater in *N*-methylpyrrolidin-2-one (NMP) than in Me_2SO, despite an equal or greater ability of NMP to solvate cations. Here the lesser ability of NMP to stabilize anions and/or its lower dielectric are likely factors leading to stronger ion pairing.[30]

Cation Effects on Ion Pairing. The factors involved in ion-pair formation are complex. In addition to the interaction between the two ions, the solvation of the cation, anion, and ion pair are important. As an illustrative example of these factors let us consider two very different organic anions, the very delocalized hydrocarbon fluorenyl anion and the localized *tert*-butoxide anion. The association constants have not been measured in the same solvent because the latter is so much more prone to ion-pair formation that it would be very difficult. Hogen-Esch and Smid have measured the association constants for fluorenyl anion with alkali metal cations in tetrahydrofuran and found the order Cs > K > Na > Li.[22] They attribute this order to the better solvation of smaller cations by THF. Since the charge in the anion is so delocalized, the solvent–cation interaction determines the order. The oppo-

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(27) Singer, D., unpublished results.

(28) Bordwell, F. G.; Algrim, D.; Fried, H. E. *J. Chem. Soc., Perkin Trans. 2* 1979, 726.

(29) Olmstead, W. N.; Margolin, Z.; Bordwell, F. G. *J. Org. Chem.* 1980, first paper in a series of six in this issue.

(30) Bordwell, F. G.; Branca, J. C.; Hughes, D. L.; Olmstead, W. N. *J. Org. Chem.* 1980, third paper in a series of six in this issue.

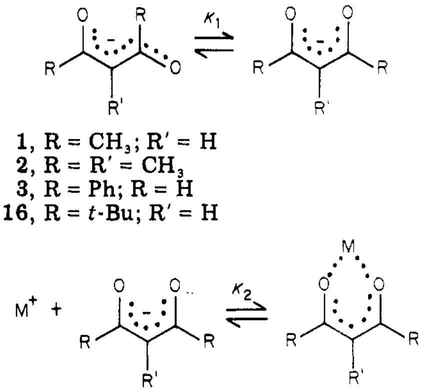
site result (Li > Na > K > Cs) was observed by Exner and Steiner for *tert*-butoxide anion in Me_2SO, despite the much stronger solvation of the smaller cations.[23,31] In this case the anion is localized, and the stronger interaction between it and the smaller cations determines the order of alkali metal association constants.

For the β-dicarbonyl compounds in this and other studies[6,7] the ion-pairing interactions decrease with increasing cation size, just as for *tert*-butoxide anion. This order is determined by the cation–anion interaction, which is strongest for lithium in the chelated ion pair. A different cation order (Na > Li > K) is observed for the methanenitronate ion (**7**) and the benzoate ion (**8**). Although the reversal is modest, it is striking when compared to the β-keto enolate ion association constants. The log K_{as} values for **7** and **8** are larger with K^+ (1.7) than for the aceto enolate ion **1** (1.4); those for Na^+ are of comparable size for all three anions (2.7, 2.3, and 2.6, respectively). On the other hand those for **7** and **8** with Li^+ (2.25 and 2.2) are much smaller than for **1** (4.8). The data suggest that for **7** and **8**, Li^+ does not coordinate specifically with both oxygen atoms, as it presumably does in the six-membered-ring keto enolate ion complexes. It seems likely that the relatively small oxygen–oxygen distances found in the planar $PhCO_2^-$ and $H_2C{=}NO_2^-$ ions is not sufficient to permit optimal chelation with Li^+ cation but that the larger Na^+ and K^+ cations can still be accommodated reasonably well.

This anomalous alkali metal ion stability order does not occur for several carboxylic acids in two other solvents. Thus, for trifluoroacetate ion in propylene carbonate the order of increasing ion-pair stability is Cs < Rb < K < Na < Li.[32a] Likewise, the order for a number of substituted benzoic acids in 2-butanone is K < Na < Li.[32b] The reason behind the different order in Me_2SO is that smaller cations are solvated more strongly than larger cations relative to the other two solvents. For instance, the free energies of transfer from propylene carbonate to Me_2SO for Li^+, Na^+, and K^+ are −9.2, −6.9, and −4.3 kcal/mol.[24] The greater solvation of Li^+ leads to smaller Li^+–anion interaction. These differing alkali metal ion orders illustrate the important influence of solvation on ion pairing (vide supra).

Conformational Effects on Ion Pairing. As discussed above for the acidities of these organic acids, the conformations of the β-dicarbonyl anions are important factors in determining the degree of ion-pair formation. The simplest and most dramatic example of this is the difference in behavior of **5** and **6**. Anion **5** is constrained to be in the *Z,Z* or *U* conformation, which is the best for chelation of the cation. The association constants for this anion are substantially larger than those for any other anion in Table II. On the other hand, **6** cannot adopt the *U* conformation, which is reflected in its low association constant.

Scheme I



1, R = CH_3; R′ = H
2, R = R′ = CH_3
3, R = Ph; R = H
16, R = *t*-Bu; R′ = H

$$K_{as} = \frac{K_1K_2}{1 + K_1}$$

The open-chain β-dicarbonyl anions can exist in three conformations, as shown in **9**–**11**. Raban et al. have shown that the free anions have two different NMR methyl signals.[20] Therefore they exist in the *E,Z* conformation. Addition of an alkali metal salt forces the anions into the *Z,Z* conformation, which can also be identified in the NMR. Thus there are two equilibria, shown in Scheme I, which are included in the measured association constants for all of the acyclic β-dicarbonyl anions. For **1**, K_1 is small ($\ll$ 1, from the NMR work of Raban[20]). For **3**, K_1 is larger because of the steric repulsion between the oxygen and phenyl ring in the *E,Z* form. This is reflected in the higher association constant. Similar behavior occurs in the anion for **16**, where R = *t*-Bu and R′ = H; in the latter case log K_{as} = 3.12 for K^+ and 3.47 for Na^+.[7a] For **2** (R = R′ = CH_3), there is significantly less ion pairing than for acetylacetonate (**1**). This is almost certainly a conformational effect also. Apparently K_1 for **2** is even smaller than that for **1**. The origin of this effect is unclear. A similar difference has been observed for the anions from ethyl acetoacetate[6b] and ethyl ethylacetoacetate in hexamethylphosphoramide.[6c]

The observation of triple ions (LiA_2^-) for **5** (Table II) is not novel. Raban has similar evidence for **1** and Li^+ in methanol.[20] Presumably at higher concentrations all of the species in Table II will form triple ions and higher aggregates.

There are many factors involved in ion-pairing phenomena. Because of their great structural variety, organic anions are ideal for delineating some of these factors. The method presented here for quantitatively measuring the amount of association of organic electrolytes in solution will be useful in the study of these factors, as demonstrated here for conformations of β-dicarbonyl anions and the anion geometries of methanenitronate and benzoate. In addition, the method is so convenient that it can be used to address questions that occur in specific examples of ion-pairing effects on organic reactions.

Experimental Section

The details of the experimental procedures for drying Me_2SO, synthesizing $K^+CH_3SOCH_2^-$, and making the spectrophotometric measurements appear elsewhere.[9,29] Following addition of the acid to the Me_2SO solution of indicator acid and anion, as described in those two references, several (five to ten) weighed aliquots of an Me_2SO solution of the appropriate alkali metal iodide were added to the spectrophotometric cell, the absorbance being measured after each addition. The most consistent results were obtained when the percentage of ion pairing, as defined in Table I, was kept between 10 and 90. Therefore, the concentrations of acid anion and metal cation were varied, depending on the magnitude of K_{as}. For the acids presented here the ion concentrations were between 10^{-5} and 2×10^{-2} M.

Solutions of $Na^+CH_3SOCH_2^-$ were prepared in a manner similar to that of the potassium salt from NaH (Alfa). The reaction is considerably slower, taking several hours at room temperature under vacuum. This procedure was found to give a better grade of base than the originally published method of heating the solution to 65–70 °C under nitrogen.[33] All of the compounds except

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(31) The free energies of solvation of K^+, Na^+, and Li^+ in Me_2SO are −83.5, −101.5, and −125.6 kcal/mol. Data are from ref 2, p 192.

(32) (a) Jansen, M. L.; Yeager, H. *J. Phys. Chem.* **1974**, *78*, 1380. (b) Crisp, S.; Hughes, S. R. C.; Price, D. H. *J. Chem. Soc. A* **1968**, 603.

(33) Corey, E. J.; Chaykovsky, M. *J. Am. Chem. Soc.* **1962**, *84*, 866; **1965**, *87*, 1345.

4-methyl-1,8-decalindione were commercially available and were purified by distillation (>99% pure by GLC) or recrystallization (pure by TLC with good melting points). Commercial samples of KI, NaI, and LiI were dried at 350 °C to constant weight. A sample of 4-methyl-1,8-decalindione[34] was provided by H. E. Zaugg and was purified to 99+% purity by GLC through distillation [bp 72.5 °C (0.1 mmHg)].

Acknowledgment. We are grateful to Dr. H. E. Zaugg for a gift of 4-methyl-1,8-decalindione, to Dr. R. J. McCallum for much help in writing the computer program for the data treatment, to Donald Algrim for the data on ion pairing of carboxamide ions, and to the donors of the Petroleum Research Fund, administered by the American Chemical Society, for support. Suggestions of Timothy S. Ungermann were helpful in interpreting some of the acidity data.

Registry No. 1 (M = K), 19393-11-4; 1 (M = Na), 15435-71-9; 1 (M = Li), 18115-70-3; 2 (M = K), 72610-66-3; 2 (M = Na), 34916-51-3; 2 (M = Li), 70902-15-7; 3 (M = K), 22658-72-6; 3 (M = Na), 19269-14-8; 3 (M = Li), 22643-60-3; 4 (M = K), 37892-24-3; 4 (M = Na), 996-82-7; 4 (M = Li), 34727-00-9; 5 (M = K), 2421-91-2; 5 (M = Na), 2421-92-3; 5 (M = Li), 2421-93-4; 6 (M = K), 37892-21-0; 6 (M = Na), 17372-26-8; 6 (M = Li), 22643-61-4; 7 (M = K), 28273-52-1; 7 (M = Na), 25854-38-0; 7 (M = Li), 53328-79-3; 8 (M = K), 582-25-2; 8 (M = Na), 532-32-1; 8 (M = Li), 553-54-8.

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(34) Zaugg, H. E.; Schaefer, A. D. *J. Am. Chem. Soc.* **1965**, *87*, 1857.

Equilibria Involving Organic Anions in Dimethyl Sulfoxide and *N*-Methylpyrrolidin-2-one: Acidities, Ion Pairing, and Hydrogen Bonding

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Received October 9, 1979

Four new indicators, 5-nitrobarbituric acid, 2,4-dinitro-1-naphthol, 4-chloro-2,6-dinitrophenol, and 2,4-dinitrophenol, have been used to complete the anchoring of the Me_2SO acidity scale. Two of these indicators plus 2-methyl-4,6-dinitrophenol were used to anchor a new acidity scale in the dipolar nonhydroxylic solvent *N*-methylpyrrolidin-2-one (NMP). Overlapping indicators and standard acids are now available to cover the pK_a range 0 to 32 in both solvents. Absolute acidities were found to be about 1 pK_a unit lower in NMP than in Me_2SO for most weak acids. This is attributed to a slightly lower basicity for NMP. For acids in which the negative charge in the anion is localized on oxygen the acidities were about 2 pK_a units lower in NMP. Homohydrogen bonding ("homoconjugation") constants for carboxylic acids and phenols were found to be about an order of magnitude larger in NMP than in Me_2SO. Ion association (ion pairing) constants were also found to be about an order of magnitude larger in NMP than in Me_2SO. The larger association constants for NMP are consistent with the more positive free energies of single-ion transfer from water to NMP than from water to Me_2SO that have been observed for small anions. Acidity constants for 48 weak acids for the pK_a range 2 to 31 in NMP are reported. Homohydrogen bonding constants for four phenols and two carboxylic acids and association constants for the anions of 10 weak acids with K^+ are given. It is suggested that solvents of the type NMP and Me_2SO be classified as "dipolar nonhydroxylic", rather than as "dipolar aprotic".

Organic chemists have been working to establish acidity scales for weak organic acids for almost 50 years to augment and complement the water acidity scale. An equilibrium acidity scale in benzene, which provided semiquantitative relative acidities for a few alcohols, ketones, hydrocarbons, and anilines, has been in use for most of this period.[1] More recently a quantitative scale of equilibrium acidities in cyclohexylamine (CHA) has provided relative acidities for many types of hydrocarbons and for some related compounds.[2] Although these scales provide important information concerning the relative acidities of groups of closely related compounds, the measured acidities are cation dependent,[3] and relative acidities may vary greatly with the medium if the structures of the acids being compared are not closely related.[4] Furthermore, since the "pK_a's" determined in such solvents are related to arbitrary reference standards in other solvents,[5] meaningful comparisons with pK_a's determined in any other medium cannot be made. For a time it appeared that a unique quantitative pK_a scale could be developed in mixed aqueous media using the H_- technique.[6] This hope was dashed, however, by the observation that compounds of different structural types showed divergent behavior in such mixtures.[7,8] Strongly dipolar nonhydroxylic ("aprotic") solvents, such as dimethyl sulfoxide (Me_2SO), dimethylformamide (DMF), acetonitrile, and hexamethylphosphoramide (HMPA), finally provided media in which equilibrium acidities directly comparable to those in water could be obtained.[9] Common oxygen acids such as carboxylic acids, phenols, and alcohols have been found to be less acidic in these media than in water by many

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[1] McEwen, W. K. *J. Am. Chem. Soc.* **1936**, *58*, 1124–1129.

[2] For paper 59 in this series see: Streitwieser, A.; Berk, C. M.; Robbers, K. *J. Am. Chem. Soc.* **1978**, *100*, 8271–8272.

[3] Streitwieser, A.; Brauman, J. I.; Hammons, J. H.; Pudjaatmaka, A. H. *J. Am. Chem. Soc.* **1965**, *87*, 384–386.

[4] Streitwieser, A.; Reuben, D. M. E. *J. Am. Chem. Soc.* **1971**, *93*, 1794–1795.

[5] The acidity scale in benzene is based on the pK_a of MeOH in MeOH,[1] that in cyclohexylamine is based on the pK_a of 9-phenylfluorene in aqueous sulfolane determined by the H_- method.[2]

[6] Stewart, R.; O'Donnell, J. P. *J. Am. Chem. Soc.* **1962**, *84*, 493–494.

[7] Steiner, E. C.; Starkey, J. D. *J. Am. Chem. Soc.* **1967**, *89*, 2751–2752. Steiner, E. C.; Gilbert, J. M. *Ibid.* **1965**, *87*, 382–384.

[8] Ritchie, C. D.; Uschold, R. E. *J. Am. Chem. Soc.* **1967**, *89*, 2752–2753.

[9] Parker, A. J. *Chem. Rev.* **1969**, *69*, 1–32. Ritchie, C. D. "Solute-Solvent Interactions"; Coetzee, J. F., Ritchie, C. D., Eds.; Marcel Dekker: New York, 1969; Vol. 1, Chapter 4, pp 219–245.