Heats of Formation of a Meisenheimer Complex and Heats of Transfer in Methanolic Dimethyl Sulfoxide Solutions

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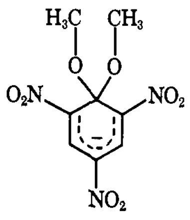
Abstract: The heat of reaction between sodium methoxide and 2,4,6-trinitroanisole to give sodium 1,1-dimethoxy-2,4,6-trinitrocyclohexadienide has been measured in methanolic dimethyl sulfoxide solutions and was found to be 15 kcal/mol more exothermic in 95% DMSO than in methanol. Heats of transfer of sodium methoxide, the Meisenheimer complex, sodium picrate, and 2,4,6-trinitroanisole from methanol to methanolic dimethyl sulfoxide solutions have been measured. The Meisenheimer complex is stabilized much more in DMSO than is the picrate anion. The transfer of sodium methoxide to DMSO rich solvents is endothermic and is a linear function of the H_- acidity function.

There has been considerable recent interest in the chemistry of Meisenheimer complexes.[2-7] Many of these investigations have involved spectroscopic studies of the structures of the complexes and the kinetics of their formation and decomposition. However, very few thermodynamic data concerning the formation of these complexes are available.[8-10] The great stabilizing effect exerted on these complexes by dipolar aprotic solvents is well known.[2-7] While much attention has been given to the heats of transfer of inorganic species from protic to dipolar aprotic solvents,[11-15] the transfer of complex organic ions and heats of transfer to mixtures of protic with dipolar aprotic solvents have been largely ignored. Much of the data on solvation by dipolar aprotic solvents has been discussed by Parker.[16] To provide quantitative information on the effect of dipolar aprotic solvents on Meisenheimer complex formation, the heat of reaction between 2,4,6-trinitroanisole (TNA) and sodium methoxide to give complex I has been measured in methanolic dimethyl sulfoxide solutions. Also, the heats of transfer of the complex I, sodium picrate, sodium methoxide, and 2,4,6-trinitroanisole across the methanol–DMSO solvent system have been determined in order to provide information about solvation in these media.



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Experimental Section

The calorimeter and its use have been described previously.[10] Methanol was dried by treating it with magnesium turnings and refluxing for 24 hr followed by careful distillation and storage over 3A molecular sieves. It contained less than 0.01% (v/v) of water as determined by gas chromatography (0.25 in. × 8 ft Porapak Q column). Fisher reagent grade DMSO was dried over 4A molecular sieves and also contained less than 0.01% water. Sodium picrate and 2,4,6-trinitroanisole were commercial preparations recrystallized from methanol and dried under vacuum over P_2O_5. We are indebted to Dr. E. J. Fendler for a gift of the purified Meisenheimer complex. The heat of solution of sodium methoxide is very sensitive to the purity of the salt. Commercial samples commonly give heats of solution several kilocalories per mole less than the carefully purified material. Sodium methoxide was prepared by treating carefully cleaned sodium with anhydrous methanol in a nitrogen atmosphere, filtering, removing the methanol under vacuum, and drying the salt at 100° under vacuum for several days. Sodium methoxide of good quality could be salvaged from commercially available samples by treating the salt with enough anhydrous methanol to dissolve about 50% of the sample, filtering the slurry in a nitrogen atmosphere followed by removal of the methanol, and drying as before. Sodium methoxide was stored and handled in a drybox or glove bag under nitrogen dried with P_2O_5. After about 1 week's storage and several exposures to the glove bag atmosphere, the sodium methoxide used in a glove bag gave a higher (less exothermic) heat of solution than freshly prepared material. This effect was not observed with material stored in a desiccator and transferred in a drybox. The values reported here were gathered over a period of 1.5 yr using several different batches of purified sodium methoxide and solvents.

Commercial sodium methoxide was purified as described above and the heat of solution in methanol was measured. The heat of solution of the unpurified material was *ca.* −12 kcal/mol. After three treatments with methanol it reached a maximum exothermic value of −17.2 kcal/mol. Two more repetitions of the purification process had no effect on the heat of solution. Sodium methoxide prepared from metallic sodium as described above gave the same heat of solution in methanol as did the purified material. Purification of this sodium methoxide using the technique described gave no change in the heat of solution. All samples of sodium methoxide prepared were periodically checked for purity by titrating with

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Table I. Heats of Solution and Reaction in MeOH–DMSO Solutions at 25°

% DMSO (v/v)	$\Delta\bar{H}_S$, kcal/mol ($NaOCH_3$)	$\Delta\bar{H}_S$, kcal/mol (TNA)	$\Delta\bar{H}_S$, kcal/mol (sodium picrate)	$\Delta\bar{H}_S$, kcal/mol (complex I)	ΔH_R (TNA + $NaOCH_3$)
0	-17.2 ± 0.4^a	$+5.60 \pm 0.27$	-1.77 ± 0.06	$+2.32 \pm 0.20$	-4.86 ± 0.30
10	-16.6 ± 0.1^b	$+4.69 \pm 0.08$	-2.65 ± 0.14		-6.41 ± 0.17
20	-15.9 ± 0.3	$+4.50 \pm 0.18$			-6.98 ± 0.22
30	-15.0 ± 0.3	$+4.36 \pm 0.23$	-3.18 ± 0.28		-8.40 ± 0.26
50	-13.3 ± 0.7	$+4.16 \pm 0.13$	-3.04 ± 0.12	-1.07 ± 0.11	-10.3 ± 0.4
60					-11.8 ± 0.2
70	-10.2 ± 0.2	$+3.88 \pm 0.09$			
80	-8.96 ± 0.32	$+3.67 \pm 0.15$	-3.23 ± 0.03	-1.79 ± 0.42	-15.3
95.4	-6.56 ± 0.32	$+3.79 \pm 0.22$		-4.46 ± 0.79	-20.5
100			-3.41 ± 0.08		

[a] Standard deviation (σ). All values are the average of at least six measurements. [b] Obtained by linear least-squares extrapolation to infinite dilution. The errors reported are the standard deviation of the intercept.

perchloric acid. Samples not 100% pure within experimental error were discarded.

Results

The heats of solution ($\Delta\bar{H}_S$) of sodium methoxide,[17] 2,4,6-trinitroanisole (TNA), the Meisenheimer complex (I), and sodium picrate, as well as the heat of reaction between sodium methoxide and 2,4,6-trinitroanisole to give the complex I, have been measured across the methanol–DMSO solvent system. All values are the average of at least six individual determinations and showed no dependence on the concentration of the salt at the low levels used in this work (10^{-4}–10^{-2} *M*) except for sodium picrate. Further studies of this salt are in progress and discussion of the concentration dependence will be deferred until more data are available. These data are shown in Table I. Above 70% DMSO, the reaction of 2,4,6-trinitroanisole with methoxide ion yields either the 1,3 Meisenheimer complex or another complex.[22] The heats of formation of these complexes as well as the kinetics of formation of the second complex have been measured and will be the subject of a subsequent report. Because of this, the heat of reaction to give the 1,1 complex (I) had to be determined indirectly using the enthalpy cycle shown in Scheme I. This enthalpy cycle was also used to obtain the heats of transfer of the Meisenheimer complex (I) below 60% methanol (Figure 1). The heat of reaction to form the 1,1 complex in solvent B ($\Delta H_{R,B}$) is given by

$$\Delta H_{R,B} = \Delta\bar{H}_{S,NaOMe,A} + \Delta\bar{H}_{S,TNA,A} + \Delta\bar{H}_{R,A} + \Delta\bar{H}_{S,Comp,B} - \Delta\bar{H}_{S,Comp,A} - \Delta\bar{H}_{S,NaOMe,B} - \Delta\bar{H}_{S,TNA,B}$$

Since all the quantities on the right can be measured, the heat of reaction can be calculated in systems where the 1,1 complex is not actually formed, but where it is stable.

Scheme I

$$\begin{tikzcd}
NaOCH_{3\,Soln.A} & + & TNA_{Soln.A} \arrow[r, "\Delta H_{R,A}"] & Complex_{Soln.A} \arrow[dd, bend left=60, "\Delta H_{T,complex}"] \\
NaOCH_{3\,Solid} \arrow[u, "\Delta\bar{H}_{S,NaOMe,A}"'] \arrow[d, "\Delta\bar{H}_{S,NaOMe,B}"'] & & TNA_{Solid} \arrow[u, "\Delta\bar{H}_{S,TNA,A}"'] \arrow[d, "\Delta\bar{H}_{S,TNA,B}"'] & Complex_{Solid} \arrow[u, "\Delta\bar{H}_{S,Comp.A}"'] \arrow[d, "\Delta\bar{H}_{S,Comp.,B}"'] \\
NaOCH_{3\,Soln.B} & + & TNA_{Soln.B} \arrow[r, "\Delta H_{R,B}"] & Complex_{Soln.B}
\end{tikzcd}$$

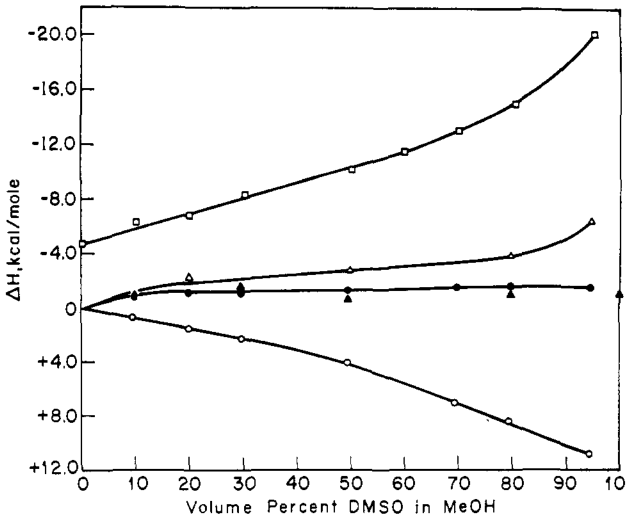

Figure 1. The heat of transfer of 2,4,6-trinitroanisole (●), sodium methoxide (○), the Meisenheimer complex I (Δ), and sodium picrate (▲) from methanol to methanol–DMSO mixtures and the heat of reaction (□) between TNA and NaOMe to give the complex I.

The heat of transfer of a species from one solvent to another is one of the most useful quantities in evaluating the differences in solvation of the species in a series of solvents. Heats of transfer from methanol to methanolic DMSO solutions for the Meisenheimer complex (I), sodium picrate, sodium methoxide, and trinitroanisole are shown in Figure 1. These plots were derived from the data in Table I.

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(17) Our value for the heat of solution of sodium methoxide in methanol (-17.2 ± 0.4 kcal/mol) disagrees with that reported by Paul[18,18a] (-7.60 kcal/mol). If the sodium methoxide used in our measurements is weighed in air, rather than in a dry nitrogen atmosphere, we obtain values ranging from -6.5 to -8.5 kcal/mol. One sample exposed to the air for only a few minutes was only 82% pure by titration. The standard deviation in any one series of measurements is often low, *ca.* 0.2–0.5 kcal/mol. The heat of neutralization of sodium methoxide reported by Paul[18] (-9.31 kcal/mol) does not agree with the data of Wolfenden[19] (-11.77 kcal/mol). In addition the heat of solution of water in methanol reported by Paul[18] (-6.7 kcal/mol) is quite different from that observed here (-0.73 ± 0.01 kcal/mol) and in other laboratories (-0.724 ± 0.012)[20] and is about the same as the heat of solution of water in 91% H_2SO_4.[21] Because of these discrepancies and because we can obtain results comparable to Paul's using wet sodium methoxide, we believe the value reported here is more accurate.

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(18a) NOTE ADDED IN PROOF. Dr. Paul has informed us that he obtained a significantly lower (more exothermic) value for the heat of solution of sodium methoxide in methanol than originally reported[18] by repeating his measurements under very dry conditions. However, there still exists a difference between his new results and ours.

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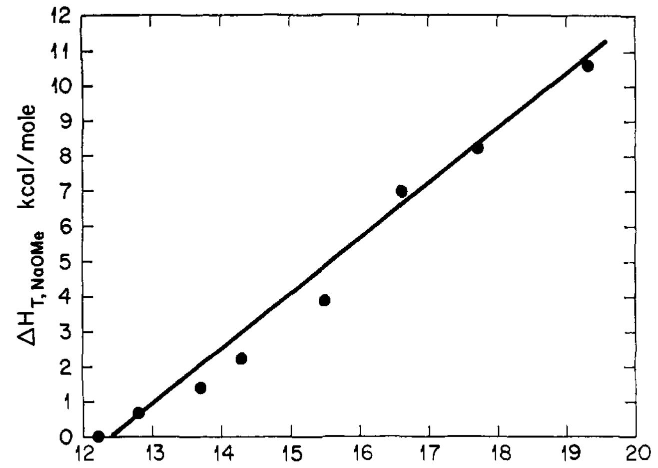

Figure 2. H_- *vs.* the heat of transfer of sodium methoxide from methanol to methanol–DMSO solutions at 25°.

Discussion

The most striking feature of the data presented in Figure 1 is the change in the heat of reaction between sodium methoxide and 2,4,6-trinitroanisole to give the complex I. As the solvent is changed from pure methanol to DMSO:MeOH 95:5 (v:v), this reaction becomes more exothermic by 15 kcal/mol. The increased stabilization of Meisenheimer complexes in dipolar aprotic solvents is a well-known phenomenon.[6,8,16,23] In this case, the chief contribution to the increased heat of reaction is the decreasing stability of sodium methoxide as the concentration of DMSO increases. The heat of transfer of sodium methoxide from methanol to 95% DMSO is endothermic by 10.6 kcal/mol. The increasing stability of the complex contributes about 7 kcal/mol to the increased heat of reaction. It is quite apparent that the enormous increase in ease of formation of Meisenheimer complexes in DMSO is due not only to the increased stability of the complexes formed, but also to the decreasing stability of the nucleophile. In this case, the latter factor is the more important one.

There are several other large solvent effects shown in Figure 1 which should be considered. When comparing two salts, sodium ion need not be considered since it is common to all of the salts. Also, since all of the heats reported are concentration independent between *ca.* 10^{-4} and 10^{-2} *M* or have been extrapolated to zero concentration, ion-pairing effects are probably not important. Compared to the large organic anions, methoxide ion is quite small and probably derives significant stabilization due to hydrogen bond formation with the solvent. This factor will become less important as the amount of methanol in the solvent decreases leading to a decrease in the stability of the methoxide ion.[8,16,24]

The increase in the stability of the Meisenheimer complex is striking, particularly when it is compared with the behavior of sodium picrate. The heats of transfer of sodium picrate and trinitroanisole are quite similar. D'Aprano and Fuoss[25] have argued convincingly that the picrate anion in acetonitrile–alcohol mixtures is solvated by hydrogen bonding of three molecules of alcohol to the negative oxygens on the three nitro groups. If this is the controlling factor (dispersion interactions should be important and different for the two species[26] and there will be a contribution due to changes in the solvation of Na^+) similar behavior for methyl picrate (TNA) and sodium picrate is expected. Also, it has been argued that the stabilization of the methoxide ion due to H-bonding decreases as the amount of DMSO increases. It is not apparent why this would not also be the case with sodium picrate, although the data of Fuoss in acetonitrile indicate that "solvent sorting" does occur extensively. In short, there should be several effects operating which could very well be cancelling each other out. It seems best to defer further speculation until more data are available.

The picrate anion and Meisenheimer complex are of roughly similar size, and the dielectric constant of the medium changes little, so it seems unlikely that the difference in their heats of transfer can be explained on the basis of simple electrostatics. Furthermore, it seems unlikely that the complex would be a significantly better or worse hydrogen bond acceptor than the picrate anion. This suggests that dipole–dipole and London dispersion forces may be controlling the situation. As Grunwald has pointed out, this is the sort of system in which dispersion interactions could be large.[26] The dipole moment of DMSO (3.90)[27] is considerably greater than that of methanol (1.70).[27] Thus the stability of the ion containing better centers of dispersion should increase in DMSO. These interactions will be a function of the oscillator strengths of the ions. Grunwald[26] has calculated these for the picrate anion and obtained 0.23 for the *p*-nitro group and 0.04 per *o*-nitro group. Using Gold's data,[28] the oscillator strength of the 4100 and 4800 Å transitions are roughly 0.20 and 0.17 l. mol^{-1} cm^{-2}, respectively. It is apparent that the oscillator strength of the complex is greater than that of the picrate ion so the complex should be stabilized in DMSO relative to the picrate anion by these interactions. A meaningful quantitative calculation of these interactions does not seem possible, but there are ample data available demonstrating that the size of this effect can be sufficient to explain the observed difference.[26,29,30] Dipole–dipole interactions are probably also significant. Investigations of the thermodynamics of transfer of other large, highly polarizable organic ions from protic to dipolar aprotic solvents are continuing in these laboratories.

Much of the interest in dipolar aprotic solvents stems from the extremely strong bases which can be generated in them. For solutions of sodium methoxide in DMSO an acidity function (H_-) has been developed.[24] A plot of H_- *vs.* the heat of transfer of sodium methoxide is shown in Figure 2. The reasonably good straight line relationship has implications concerning the activities of several species involved.

For the dissociation of an acid HA

$$H_- = -\log \frac{a_{H^+} f_{A^-}}{f_{HA}} \tag{1}$$

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where a denotes activity and f is an activity coefficient. Let us now consider the free energy of transfer of A^- ($\Delta G_{T,A^-}$) from one solvent to another and follow Hepler[31] in expressing the free energy in terms of the external and internal enthalpy and entropy, *i.e.*

$$\Delta G_{A^-} = \Delta H_{E,A^-} + \Delta H_{I,A^-} - T\Delta S_{E,A^-} - T\Delta S_{I,A^-} \quad (2)$$

The internal enthalpy and entropy will be independent of the solvent and the external quantities can be related by the isokinetic relationship ($\Delta H_E = \beta \Delta S_E$).[31] Thus the free energy of transfer of A^- becomes

$$\Delta G_{T,A^-} = RT \ln \frac{f_{A^-}'}{f_{A^-}} = \Delta H_{E,A^-}(1 - T/\beta) \quad (3)$$

The change in the acidity function as the solvent is changed is given by

$$\delta H_- = \log \frac{a_{H^+}' f_{AH}}{a_{H^+} f_{AH}'} + \log \frac{f_{A^-}'}{f_{A^-}} \quad (4)$$

Combining eq 3 and 4 gives

$$\delta H_- = \log \frac{a_{H^+}' f_{AH}}{a_{H^+} f_{AH}'} + \frac{\Delta H_{E,A^-}}{2.3RT}(1 - T/\beta) \quad (5)$$

The heat of transfer of sodium methoxide can be expressed as $\Delta H_{T,NaOMe} = \Delta H_{T,Na^+} + \Delta H_{T,^-OMe}$. Combining this with eq 5 and remembering that

$$\frac{\Delta H_{T,Na^+}}{RT}(1 - T/\beta) = -\ln \frac{a_{Na^+}}{a_{Na^+}'}$$

gives

$$\delta H_- = \log \frac{a_{H^+}' a_{Na^+} f_{MeOH}}{a_{H^+} a_{Na^+}' f_{MeOH}'} + \frac{\Delta H_{T,NaOMe}}{2.3RT}(1 - T/\beta) \quad (6)$$

Thus, if the activity coefficient term in eq 6 is a constant a straight line plot of H_- *vs.* $\Delta H_{T,NaOMe}$ is expected. If the activity coefficient ratio f_{MeOH}/f_{MeOH}' is constant, or nearly so (a rather unlikely occurrence), the observed behavior requires that the activities of H^+ and Na^+ be affected in the same way by changes in the solvent composition. This is surprising, since significant differences even in the solvation of Li^+ and Na^+ in DMSO have been observed.[32] It does not seem reasonable to expect the activity coefficient term would be a constant and this relationship is being explored further.

Acknowledgment. Partial support of this work through grants made by the Research Corporation and the National Science Foundation (to J. W. L.) as well as by the Atomic Energy Commission (to J. H. F.) is gratefully acknowledged

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Heterocyclic Radical Ions. V. Electron Spin Resonance Splittings of Semifuraquinones[1]

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Abstract: The esr spectra of several rigid, strained dialkylmaleic anhydride anions (semifuraquinones) are reported and discussed. These were prepared by reduction of Diels–Alder adducts of dichloromaleic anhydride and cyclic dienes. The polarography of several of the adducts and related compounds was investigated, and it was concluded that both C–Cl bonds are being broken in the slow step of the reduction. The adducts of cyclopentadiene and dimethylfulvene with dibromomaleic anhydride were also prepared and reduced, and esr spectra of the semifuraquinones are reported.

Long-range proton esr splitting constants are especially large in rigid, strained systems, and the factors leading to the enhanced odd-electron delocalization in such situations have been a subject of interest for about 5 years. To study such splittings, some group which can be converted to a conveniently stable radical is needed to "spin label" the aliphatic portion of the molecule of interest. Semidone spin labels have been extensively employed by Russell's group,[4] and semiquinones by Stock's group.[5] We have chosen the dialkylmaleic anhydride anion ("semifuraquinone" group) for similar studies.[1] Semifuraquinone is a very similar spin label to semiquinone in that the signs of the p orbitals in position to overlap with the aliphatic portion of the molecule (the C_α p orbitals) are opposite, but have a higher spin density in the semifuraquinone; since resolution is a major problem in these experiments, the greater expected splitting constants and the lack of splittings introduced by the label are valuable assets of semifuraquinones. Semidione is fundamentally different from the quinone-type spin labels, for the signs of the C_α p orbitals are the same; the π spin density at the C_α carbons is also somewhat higher for semidione than for semifuraquinone. The semifuraquinone spin label was also attractive to us because it is easy to introduce by Diels–Alder reactions of acetylenedicarboxylic acid

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(2) Alfred P. Sloan Fellow, 1968–1970.

(3) Ford Foundation Fellow.

(4) G. A. Russell, G. W. Holland, and K. Y. Chang, *ibid.*, **89**, 6629 (1967), and references therein.

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