cyclohexanediamine and to Drs. D. A. Buckingham and A. M. Sargeson for samples of *trans*-*R*,*R*-[Co(trien)-Cl_2]ClO_4 and the *trans*-*S*,*S* compounds, [Co(*N*-Meen)$_2$-Cl_2]ClO_4, [Co(*N*-Meen)$_2$(NO_2)$_2$]ClO_4, and [Co(*N*-Meen)$_2$-NO_2Cl]ClO_4. We also thank Dr. N. C. Payne for calculating on his program the atomic coordinates of a number of crystal structures; this information provided the background for our assertions about the sector occupancy of various atoms in the present series of complexes.

The Effects of Polar Aprotic Solvents on Linear Free-Energy Relationships in Hydrogen-Bonded Complex Formation[1]

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Abstract: A study of the effects of polar aprotic solvents on linear free-energy relationships in hydrogen-bonded complex formation has been made. The behavior of strong pK_A bases has suggested the utility of the concept of solvent-induced (partial) ionization in hydrogen-bonded complexes as distinct from partial hydrogen-bonded ion-pair formation or ionic dissociation.

In recent work,[2] linear free-energy relationships in hydrogen-bonded complex formation were reported which follow eq 1.

$$\log K_f = m(\mathrm{p}K_{HB}) + c \qquad (1)$$

K_f is the formation constant of the 1:1 hydrogen-bonded complex for a given proton donor with a series of bases in a given solvent (generally CCl_4) at a given temperature. The pK_{HB} value is the logarithm of the corresponding formation constant for the *p*-FC_6H_4OH hydrogen-bonded complex in carbon tetrachloride at 25°. The constants, *m* and *c*, are characteristic of the proton donor, the temperature, and the solvent. Increasing values of these constants denote increased stability (decreased standard free energy) of a series of hydrogen-bonded complexes.

It has been established by other investigators that increasingly polar aprotic solvents act to decrease K_f values for a given acid–base complex compared to the value in CCl_4.[3] The present work was undertaken to provide a critical test of the application of eq 1 to the effects of aprotic solvents on the formation of the hydrogen-bonded complexes. The pure aprotic solvents cyclohexane, chlorobenzene, *o*-dichlorobenzene, 1,2-dichloroethane, and methylene chloride were selected for study. The use of *p*-FC_6H_4OH as the reference acid makes possible the application of the F nmr method developed previously[4] for determination both of the formation constants, K_f, and the F nmr shift, Δ. The latter quantity probes the electron density at the atomic center of complex formation. This method is readily applicable to these solvents, whereas the familiar ir method for determination of K_f values is not applicable because of the interference of the solvent C–H stretching frequencies.

The bases selected for study are: dioxane, benzonitrile, 3-bromopyridine, cyclohexanone, cyclopropylamine, triethylamine, quinuclidine, pyridine, diphenyl sulfoxide, *N*,*N*-dimethylformamide (DMF), dimethyl sulfoxide (DMSO), quinuclidine, 4-dimethylaminopyridine, triphenylphosphine oxide, and hexamethylphosphoramide (HMPA). This series of bases covers a very substantial range of pK_{HB} values within which there are marked variations in the nature of the atomic center of complexing and in the polar, steric and resonance effects acting upon this center. The series includes members for which pK_{HB} values parallel aqueous pK_A values (*e.g.*, the series of 3- and 4-substituted pyridines), as well as members for which there are widely diverse relationships between pK_{HB} and corresponding aqueous pK_A values[2,4] (*e.g.*, the series dioxane, benzonitrile, cyclohexanone, triethylamine, pyridine, dimethyl sulfoxide, and hexamethylphosphoramide).

In addition to eq 1, two other relationships have been investigated in this study: the linear SEE relationship[4,5]

$$\log K_f = m'\Delta + b \qquad (2)$$

where K_f and Δ values are for a given aprotic solvent, and the relationship between corresponding F nmr shift parameters, Δ, in the different aprotic solvents (S)

$$\Delta_S = m''\Delta_{CCl_4} \qquad (3)$$

Equation 2 involves the use of the *p*-FC_6H_4 tag to probe the relationship between a presumed measure of electron-density change on complex formation[4] and the corresponding standard free-energy change.

We also report in this paper the determination of new pK_{HB} and corresponding Δ(CCl_4) values for additional bases not previously reported. In several instances,

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[1] This work was supported in part by the Public Health Service, Project No. GM 14078.

[2] R. W. Taft, D. Gurka, L. Joris, P. von R. Schleyer, and J. W. Rakshys, *J. Amer. Chem. Soc.*, 91, 4801 (1969).

[3] (a) S. D. Christian, J. R. Johnson, H. E. Affsprung, and P. J. Kilpatrick, *J. Phys. Chem.*, 70, 3376 (1966); (b) H. Buchowski, J. Devaure, P. V. Huong, and J. Lascombe, *Bull. Soc. Chim. Fr.*, 2532 (1966); (c) Dr. P. V. Huong, private communication.

[4] D. Gurka and R. W. Taft, *J. Amer. Chem. Soc.*, 91, 4794 (1969).

[5] (a) R. W. Taft and L. D. McKeever, *ibid.*, 87, 2489 (1965); (b) C. S. Giam and R. W. Taft, *ibid.*, 89, 2397 (1967).

Table I. Formation Constants, K_f, and F Nmr Shifts, Δ in ppm, for Hydrogen-Bonded Complex Formation of Bases with p-C_6H_4OH at 25° in the Indicated Solvents

	Solvents														
	c-C_6H_{12}			C_6H_5Cl			o-$C_6H_4Cl_2$			$ClCH_2CH_2Cl$			CH_2Cl_2		
Base	K_f	log K_f	Δ	K_f	log K_f	Δ	K_f	log K_f	Δ	K_f	log K_f	Δ	K_f	log K_f	Δ
1. Dioxane	6.7	0.83	1.68	3.5	0.55	1.06	2.8	0.45	1.23	1.23	0.09	1.15	1.38	0.14	1.31
2. Benzonitrile	10.0	1.00	1.85	3.0	0.48	1.11	2.5	0.39	1.25	0.71	−0.15	1.15			
3. 3-Bromopyridine	12.6	1.42	2.16	11.5	1.06	1.65	8.9	0.95	1.77	5.3	0.72	1.75			
4. Cyclohexanone							10.5	1.02	1.59	3.3	0.52	1.52	3.2	0.50	1.63
5. Cyclopropylamine	15.0	1.70	2.50												
6. Triethylamine	98.0	1.99	2.82	69.0	1.84	2.41	85.0	1.93	2.50	50	1.70	2.45	47	1.67	2.52
7. Pyridine	107	2.03	2.70	40	1.60	2.18	43	1.63	2.20	19.5	1.29	2.12	18.0	1.26	2.20
8. Diphenyl sulfoxide													13.2	1.12	1.98
9. N,N-Dimethylformamide	200	2.30	2.97	55	1.74	2.13	50	1.70	2.14	18.6	1.27	2.03	15.2	1.18	2.15
10. Dimethyl sulfoxide	360	2.56	2.80	160	2.20	2.11	150	2.18	2.21	44.7	1.65	2.10	27.6	1.44	2.30
11. Quinuclidine							470	2.67	2.97						
12. N,N-Dimethylaminopyridine	1400	3.15	3.60	240	2.38	3.03	410	2.61	2.93	138	2.14	2.79	126	2.10	2.89
13. Triphenylphosphine oxide							450	2.65	2.65	110	2.04	2.55	80	1.90	2.71
14. Hexamethylphosphoramide	6300	3.8	4.1	1150	3.06	3.26	1150	3.06	3.08	355	2.55	2.92	234	2.37	3.00

Table II. New pK_{HB} and Δ Values[a]

O bases	Δ	pK_{HB}	N bases	Δ	pK_{HB}
$(C_6H_5)_3AsO$	4.27	3.66	Tetramethylguanidine	3.70	3.14
$(C_6H_5)_3PO$	3.46	3.16	Quinuclidine	2.96	2.63
$(MeO)_3PO$	2.71	2.45	N-Methylimidazole	3.20	2.50
$Et_2NC(=O)Me$		2.47	DABCO	2.75	2.20
N-Methylpyridone	3.00	2.38	4-Methoxypyridine	2.77	2.13
Tetramethylurea	3.00	2.30	N,N-Dimethylcyclohexylamine	2.71	2.08
tert-$BuC(=O)NMe_2$	2.75	2.15	N,N-Dimethylbenzylamine	2.34	1.56
$EtOC(=O)NEt_2$	2.52	1.78	Triallylamine	2.1	1.30
Acetylferrocene	2.17	1.65	3-Bromopyridine	1.99	1.30
$MeSC(=O)NEt_2$	2.48	1.56	p-Dimethylaminobenzonitrile	2.10	1.23
4-Dimethylaminobenzaldehyde	2.33	1.53	p-Methoxybenzonitrile	1.83	0.97
Xanthone		1.36	Acetonitrile	1.88	0.90
p-Methoxyacetophenone	2.12	1.33	3,5-Dichloropyridine	1.63	0.80
Cyclohexanone	2.12	1.32	Benzonitrile	1.71	0.80
2-Butanone	2.02	1.19	Tripropargylamine		0.70
Ethyl acetate	1.85	1.09	p-Bromobenzonitrile		0.57
$CF_3C(=O)NEt_2$		1.06	m-Bromobenzonitrile		0.53
$N{\equiv}CC(=O)NEt_2$	1.74	1.04	N,N-Dimethylaniline		0.45
Et_2O	1.85	1.01	Tribenzylamine		0.37
$CF_3C(=O)NMe_2$	1.97	0.89	$CF_3CH_2NEt_2$		0.23
Benzaldehyde	1.74	0.80			
Dioxane	1.45	0.73			

[a] K_f values are estimated to be accurate to ±10%. The ir method[6] has been used for each base and the K_f value obtained generally agrees with the value from the F nmr method to this order of precision (*cf.* examples given in ref 4). Δ values are estimated to be accurate to ±0.05 ppm.

revised values are given. Both the F nmr and ir methods[4,6] have been utilized.

Results

Table I lists values of the formation constant, K_f, and log K_f, for p-FC_6H_4OH obtained with the various bases in five aprotic solvents at 25.0°. Table I also gives corresponding values of the F nmr shift parameter, Δ, for complete formation of the 1:1 complex p-FC_6H_4OH. The results were obtained by previously described F nmr methods. Table II lists new values of pK_{HB} and Δ in CCl_4 solution at 25.0°.

Discussion

Plots of log K_f *vs.* pK_{HB} for the various polar aprotic solvents do not give single linear relationships (eq 1) of satisfactory precision. Only in cyclohexane (Figure 1) is such a relationship generally observed. (A small deviation for DMSO provides an exception in this medium.) In all of the polar solvents, however, eq 1 is observed to hold satisfactorily for the bases benzonitrile, dioxane, DMF, DMSO, $(C_6H_5)_3PO$, and HMPA (*i.e.*, for all bases of low pK_A). Figure 2 shows the log K_f–pK_{HB} plot for the solvent o-dichlorobenzene, which is typical. The m and c values based upon these points are given in Table III. Both m and c values decrease, in general, with increasing solvent polarity.

Table III. Summary of Limited Linear Free-Energy Equation 1 Results

Solvents	m	c
Cyclohexane	1.10	0.04
CCl_4	(1.00)	(0.00)
C_6H_5Cl	0.91	−0.16
o-$C_6H_4Cl_2$	0.90	−0.20
$Cl(CH_2)_2Cl$	0.87	−0.62
CH_2Cl_2	0.82	−0.54

The behavior of strong pK_A bases (pyridines and amines) in eq 1 is also dependent upon the solvent polarity. Thus, in the weakly polar chlorobenzene solvent, all pyridine bases (3-Br, H, and 4-NMe_2)

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(6) E. M. Arnett, L. Joris, E. Mitchell, T. S. S. R. Murty, T. M. Gorrie, and P. v. R. Schleyer, *J. Amer. Chem. Soc.*, **92**, 2365 (1970).

Table IV. Deviations (in Log K_f Units) of Amines from Equation 1

Base	pK_A	Cyclohexane	Chlorobenzene	o-$C_6H_4Cl_2$	$Cl(CH_2)_2Cl$	CH_2Cl_2
3-Bromopyridine	2.85	None	None	None	0.15	
Pyridine	5.21	None	None	0.11	0.22	0.26
4-*N*-Dimethylpyridine	9.59	None	None	0.24	0.28	0.34
Et_3N	10.72	None	0.24	0.36	0.60	0.60
Quinuclidine	11.15	None		0.42		

follow eq 1 (*cf.* Figure 3). Only triethylamine shows a small deviation. On the other hand, in the most polar solvents, 1,2-dichloroethane and methylene chloride, all pyridines and amines deviate from eq 1. The deviations from eq 1 of these strong pK_A bases are summarized in Table IV. All deviations are positive, *i.e.*, log K_f values for these bases in polar solvents are greater than expected by eq 1. Deviations clearly increase in magnitude with *increasing pK$_A$ value* of the base and with increasing polarity of the aprotic solvents. These deviations from eq 1 might be ascribed to the concurrent formation of a measurable fraction of the isomeric hydrogen-bonded ion-pair complex (BH^+---A^-).[7]

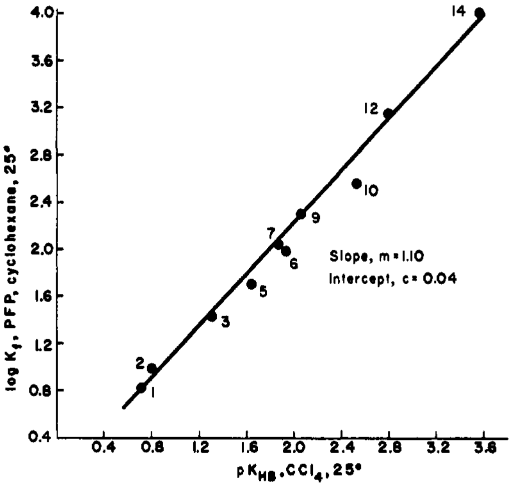

Figure 1. Linear free-energy relationship for hydrogen-bonded complex formation with p-FC_6H_3OH in nonpolar aprotic solvents: ordinate, log K_f, cyclohexane, 25°; abscissa, log K_f = pK_{HB}, CCl_4, 25°; numbers refer to bases as listed in Table I.

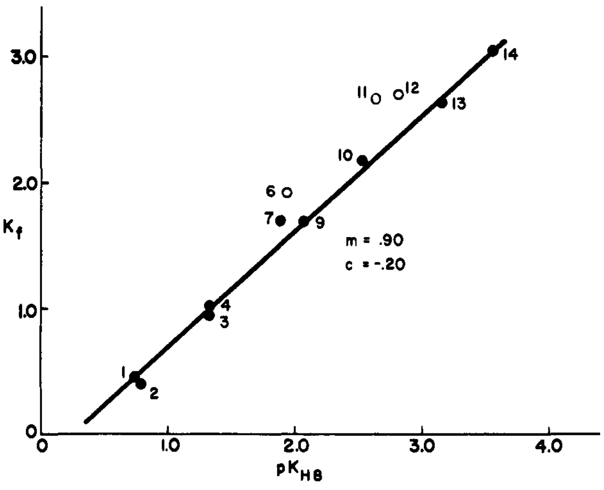

Figure 2. Linear free-energy relationship for hydrogen-bonded complex formation with p-FC_6H_4OH in polar aprotic solvent: ordinate, log K_f, o-dichlorobenzene, 25°; abscissa, log K_f ≡ pK_{HB}, CCl_4, 25°; numbers refer to bases as listed in Table I; open-circle points are amines and pyridines.

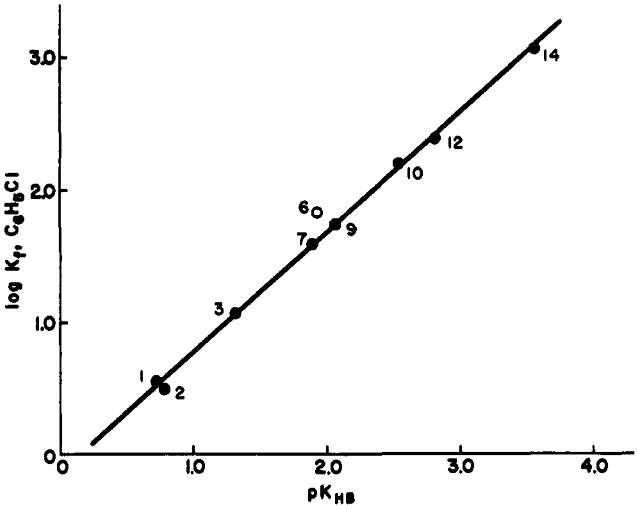

Figure 3. Linear free-energy relationship for hydrogen-bonded complex formation with p-FC_6H_4OH in polar aprotic solvent: ordinate, log K_f, chlorobenzene, 25°; abscissa, log K_f ≡ pK_{HB}, CCl_4, 25°; numbers refer to bases as listed in Table I; open-circle point is for triethylamine.

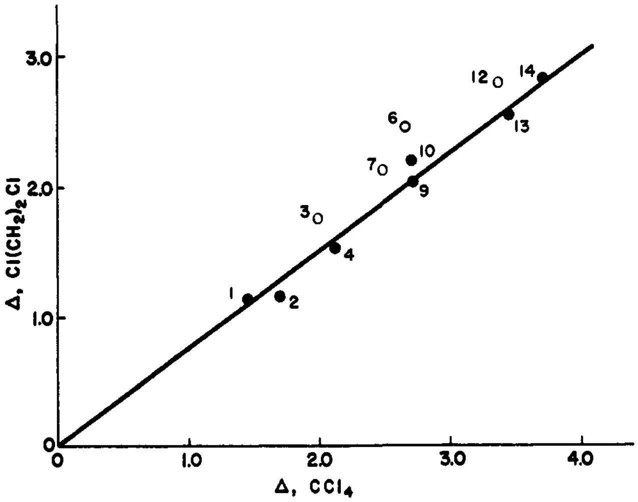

Figure 4. F nmr Δ *vs.* Δ relationship (eq 3): ordinate, Δ, 1,2-dichloroethane, 25°, ppm; abscissa, Δ, CCl_4, 25°, ppm; numbers refer to bases as listed in Table I; open-circle points are amines and pyridines.

However, we believe this interpretation to be very unlikely for the following reasons. First, the ultraviolet

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[7] (a) C. L. Bell and G. M. Barrow, *J. Chem. Phys.*, 31, 1158 (1959); (b) H. Baba, A. Matsukama, and H. Kokubun, *ibid.*, 41, 895 (1964); *Spectrochim. Acta, Sect. A*, 25, 1709 (1969).

Table V. Deviations (ppm) of Strong Bases from Equation 3

Base	Cyclohexane	Chlorobenzene	o-$C_6H_4Cl_2$	$Cl(CH_2)_2Cl$	CH_2Cl_2
3-Bromopyridine	None	0.19	0.16	0.24	
Pyridine	None	0.20	0.18	0.25	0.22
4-*N*-Dimethylpyridine	None	0.19	0.26	0.26	0.22
Et_3N	None	0.28	0.34	0.45	0.42
Quinuclidine			0.55		

spectrum of p-FC_6H_4OH in 1,2-dichloroethane solution does not show the appearance of two separate peaks with an isosbestic point upon addition of triethylamine, as does the spectrum of *p*-nitrophenol.[7b] Secondly, the F nmr Δ values show the same kind of behavior in eq 3 as do log K_f values in eq 1. That is, Δ values for the strong pK_A bases in the polar aprotic solvents give positive deviations from the linear relationships defined by very weak pK_A bases (*cf.* Figure 4 for the 1,2-dichloroethane solvent, which is typical). The deviations from eq 3 are summarized in Table V. We believe the magnitudes of the deviations in log K_f and Δ values given in Tables IV and V are not compatible with the coexistence of isomeric hydrogen-bonded complexes and hydrogen-bonded ion pairs. If the deviations in log K_f were due to contributions from hydrogen-bonded ion pairs, then the corresponding deviations in Δ values would require that the limiting shifts of the hydrogen-bonded complexes and hydrogen-bonded ion pair of a given base be the same within a few tenths of a part per million. The shift for essentially complete proton removal from p-FC_6H_4OH is 15 ppm.[8] From this figure we conclude that ion-pair formation could require a limiting shift of at least 7.5 ppm (*i.e.*, greater than 50% apparent proton transfer).[9] None of the bases give rise to shifts (observed or calculated from log K_f deviations) of this magnitude, including quinuclidine in *o*-dichlorobenzene (which gives the largest Δ deviation, Table V).

We believe the deviations in log K_f and Δ values of Tables IV and V are very probably due to an increased extent of proton transfer in hydrogen-bonded complexes permitted by the action of polar aprotic solvents. That is, polar aprotic solvents induce in the hydrogen-bonded complexes of strong pK_A bases an increased extent of ionization (stretching) of the OH bond over that anticipated by eq 1 and 3. While the resulting increase in stability (decreased free energy) is not large (less than 1 kcal/mol) it must be borne in mind that the effect on the enthalpy of complex formation is probably appreciably greater. This conclusion follows from the expected entropy loss from the solvent orientation associated with increased ionization[10] which opposes (in the free energy) the enthalpy effects.

The concept of stabilizing solvent-induced ionization in the hydrogen-bonded complexes of strong pK_A bases needs to be drawn in contradistinction to the destabilizing effects of polar aprotic solvents. The formation of the hydrogen-bonded complexes of weak pK_A bases evidently results in a loss of unbalanced electrical forces which are present in both the free acid and the free base. Thus, polar aprotic solvents interact more strongly with the free acid and base forms than with the hydrogen-bonded complex.[11] This is the "normal" (low pK_A base) behavior which is further described by eq 1 in terms of decreasing *m* and *c* values with increasing solvent polarity. Solvent-induced partial ionization provides a mechanism by which the hydrogen-bonded complexes of strong pK_A bases may more favorably compete with the free acid and base in solvation.[11] The "normal" decrease in log K_f values with increased polarity of the aprotic solvent is thus "resisted" by the solvent-induced partial ionization in these hydrogen-bonded complexes.

The behavior of log K_f *vs.* Δ plots for the various solvents is also interesting and instructive. In CCl_4, eq 2 has been shown to be followed generally.[4] The only exceptions are bases of very large steric requirements, *e.g.*, $(n\text{-}Bu)_3N$, and a few anomalous bases, *e.g.*, DMSO. The results in cyclohexane also follow eq 2 satisfactorily, with DMSO showing the only appreciable deviation. In chlorobenzene, *o*-dichlorobenzene, and 1,2-dichloroethane, eq 2 is followed satisfactorily by all of the low pK_A bases, except DMSO.

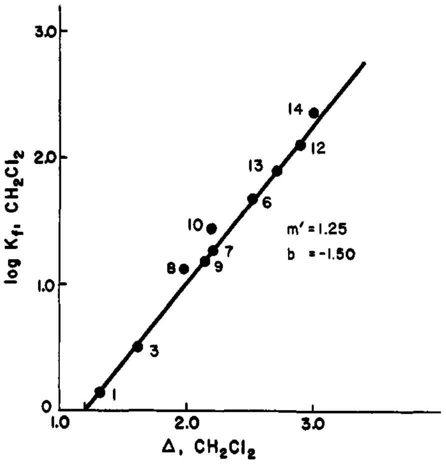

Figure 5. Linear SEE relationship for hydrogen-bonded complex formation with p-FC_6H_4OH in polar aprotic solvents: ordinate, log K_f, methylene chloride, 25°; abscissa, Δ, methylene chloride, 25°, ppm; numbers refer to bases as given in Table I.

Pyridines and amines tend to deviate from this "normal" behavior in the direction that solvent-induced partial ionization in the complex leads to relatively larger F nmr shifts than to increased log K_f values. These deviations, which are not large, are also not general to all of the polar aprotic solvents, however. In CH_2Cl_2, eq 2 holds with satisfactory precision for all bases (Figure 5). It is also of interest that the deviation of the DMSO point in the various solvents decreases in magnitude as the solvent polarity increases. It was previously suggested[4] that the smaller observed

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(8) D. Gurka, R. W. Taft, L. Joris, and P. von R. Schleyer, *J. Amer. Chem. Soc.*, **89**, 5957 (1967).

(9) *Cf.* H. B. Yang and R. W. Taft, *ibid.*, **93**, 1310 (1971).

(10) *Cf.* ref 7b.

(11) S. D. Christian, *J. Amer. Chem. Soc.*, **91**, 6514 (1969).

Table VI. Values of m' and b Parameters of Equation 2

Solvent	m'	b
$c\text{-}C_6H_{12}$	1.21	−1.21
CCl_4	1.25	−1.28
C_6H_5Cl	1.17	−0.75
$o\text{-}C_6H_4Cl_2$	1.33	−1.16
$Cl(CH_2)_2Cl$	1.43	−1.63
CH_2Cl_2	1.25	−1.50

Table VII. Values of the m'' Parameter of Equation 3

Solvent	m''
$c\text{-}C_6H_{12}$	1.08
CCl_4	(1.00)
C_6H_5Cl	1.00
$o\text{-}C_6H_4Cl_2$	0.81
$Cl(CH_2)_2Cl$	0.76
CH_2Cl_2	0.80

Δ value than that expected by eq 3 might be the result of interaction between the negative phenolic oxygen and the positive sulfur of the DMSO in the hydrogen-bonded complex. This interaction, which would increase log K_f and decrease the Δ value as observed, would also be expected to "fade away" with increased solvent polarity. Decreasing self-association of DMSO in solvents of greater polarity may be an alternate explanation for this observation. However, it is difficult with this explanation to see why the self-association should always lead to decreased apparent values of Δ and increased apparent values of log K_f. No evidence of the effects of self-association has been obtained as a concentration dependence on K_f values.[4,6]

Table VI gives the m' and b constants obtained with the use of eq 2. Table VII gives values of the m'' coefficients of eq 3. The mean value of m' for all six aprotic solvents is 1.27 ± 0.07. The individual values adhere very well to this mean. The somewhat higher value obtained in 1,2-dichloroethane is probably not significant. Values of the intercept, b, in eq 2 show no recognizable pattern of behavior. The m'' values of Table VII show the same general behavior as the m values of eq 1 (*cf.* Table III), *i.e.*, decrease with increasing solvent polarity. All of the Δ *vs.* Δ_{CCl_4} plots pass through the origin, except for the solvent chlorobenzene. In this solvent an intercept of −0.52 is obtained, which apparently reflects some constant magnetic contribution to the Δ values in this solvent. In general, however, the very similar pattern of behavior of m and m'' values, as well as the nearly constant value of m', support the use of F nmr shielding parameters as a free-energy measure (*cf.* eq 2).

Summary

Linear free-energy relationships in the form of eq 1 describe generally the effects of aprotic polar solvents on the formation of hydrogen-bonded complexes of *p*-fluorophenol with weak pK_A bases. Both m and c values in eq 1 tend to decrease with increasing solvent polarity. Similar linear relationships are found for the F nmr shifts, Δ, eq 3. The effects of aprotic polar solvents on Δ values are correlated well by the corresponding log K_f values, eq 2. The essentially constant value of m' (1.27 ± 0.07) in eq 2 indicates, as previously found, that about 1.8 kcal in the standard free-energy charge corresponds to 1.0 ppm in the F nmr shift in hydrogen-bonded complex formation with $p\text{-}C_6H_4OH$. Strong bases (amines and pyridines) are found to give positive deviations from eq 1 and 3. The deviations increase in magnitude with increasing pK_A of the base and with increasing polarity of the aprotic solvent. Solvent-induced partial ionization in the hydrogen-bonded complexes of these strong pK_A bases is considered to permit the complex to more favorably compete with the free acid and base in solvation by the aprotic polar solvent.

Hydrogen-Bonded Complex Formation with 5-Fluoroindole. Applications of the pK_{HB} Scale[1]

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Contribution from the Department of Chemistry, University of California, Irvine, California 92664. Received September 15, 1971

Abstract: The formation of the hydrogen-bonded complexes of 5-fluoroindole with 27 bases of widely differing types have been studied in CCl_4 at 25° using F nmr and ir methods. The formation constants, K_f, and F nmr shifts, Δ, for this N–H proton donor are found to correlate satisfactorily with the pK_{HB} scale, which is based upon the reference acid, $p\text{-}FC_6H_4OH$. Other generalized correlations with the pK_{HB} scale are summarized.

Linear free-energy relationships were shown recently to have broad applicability in the formation of hydrogen-bonded complexes of various OH reference acids.[2] A scale of proton acceptor strengths in hydrogen-bonded complex formation based upon *p*-fluorophenol in CCl_4 solution was defined as the pK_{HB} parameter and values of this parameter for numerous bases were tabulated. The pK_{HB} scale bears little or

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(1) This work was supported in part by the Public Health Service, Project No. GM14078.

(2) R. W. Taft, D. Gurka, L. Joris, P. von R. Schleyer, and J. W. Rakshys, *J. Amer. Chem. Soc.*, 91, 4801 (1969).