THE JOURNAL OF Organic Chemistry

VOLUME 51, NUMBER 5

MARCH 7, 1986

Equilibrium Constants for Double Hydrogen Bonding by 1,8-Biphenylenediol in Cyclohexane Solution[1]

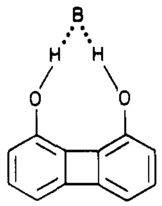
Jack Hine,* Soonkap Hahn, and David E. Miles

Department of Chemistry, The Ohio State University, Columbus, Ohio 43210

Received October 3, 1985

Equilibrium constants in cyclohexane at 23.5 °C were determined by ultraviolet measurements for hydrogen bonding of several phenols to various bases. Values for 1,8-biphenylenediol increase in the following order: 1,3-dioxolane, propionitrile, 1,4-dioxane, diethyl ether, cyclopropylamine, tetrahydrofuran, dimethylcyanamide, dimethylformamide, tetramethylurea, 1,3-dimethyl-2-imidazolidinone, dimethyl sulfoxide, tetramethylene sulfoxide, and hexamethylphosphoramide. For the oxygen bases this is the same order observed for *m*-nitrophenol and for the bases studied, *p*-nitrophenol and *p*-fluorophenol. On a per hydroxy group basis the equilibrium constants for 1,8-biphenylenediol (pK_a = 8.01 or 8.31 per hydroxy group) are somewhat smaller than those for *m*-nitrophenol (pK_a = 8.36) or *p*-nitrophenol (pK_a = 7.15) for the weaker bases, but they are much larger for the stronger bases. log–log plots of the equilibrium constants for complexing of 1,8-diol with the oxygen bases vs. the same constants for *m*-nitrophenol and *p*-nitrophenol give lines of slopes 1.37 and 1.31, respectively. The equilibrium constants for complexing of the three nitrogen bases with the 1,8-diol were too small by three- to fivefold to fit on the best line through the points for oxygen bases. These observations are interpreted in terms of double hydrogen bonding by the 1,8-diol to the oxygen bases. There appear to be significant amounts of single hydrogen bonding with the weaker oxygen bases and all the nitrogen bases. It is possible that dimethylcyanamide is doubly hydrogen bonded to a significant extent.

1,8-Biphenylenediol is a species that, in the conformation with the hydroxy groups coplanar with the rings and oriented inward, has an almost optimum geometry for forming two hydrogen bonds simultaneously to the same basic atom. X-ray crystal structure studies of adducts of the diol with hexamethylphosphoramide, 1,2,6-trimethyl-4-pyridone, and 2,6-dimethyl-γ-pyrone show that double hydrogen bonds are formed to an oxygen atom in each of these bases.[2] This establishes the existence of such structures, but it does not tell much about the stabilities of such species relative to the stabilities of the corresponding singly hydrogen-bonded adducts; in fact, it does not even show that the doubly hydrogen-bonded adducts are ever the predominant adducts in solution, where the extra internal rotation possible would favor a singly hydrogen-bonded structure. To shed more light on these points we have determined equilibrium constants for complexing of the diol and of certain phenolic model compounds to various bases in cyclohexane, where hydrogen bonding to the solvent should be minimal.



Results and Discussion

Determination of Equilibrium Constants. When increasing concentrations of a base are added to a dilute solution of a phenol in cyclohexane, the apparent extinction coefficient, ϵ_A,[3] of the phenol ordinarily changes. The bases used do not absorb appreciably at the wavelengths used, and the same amount of base was added to the reference cell as to the sample cell. With some phenol–base pairs, such as 1,8-biphenylenediol and *N,N,N′,N′,N″,N″*-hexamethylphosphoramide, as shown in Figure 1, there is a fairly linear change in ϵ_A until about 1 mol of base has been added per mol of phenol; with larger base concentrations, ϵ_A quickly becomes essentially constant. This establishes the stoichiometry of the complex. If there is enough curvature between the two linear portions of the plot, the equilibrium constant can be determined. The

---

(1) (a) This investigation was supported in part by NSF Grant CHE-8114770. (b) Abstracted in part from the Ph.D. Dissertations of David E. Miles, The Ohio State University, 1982, and of Soonkap Hahn, The Ohio State University, 1985.

(2) Hine, J.; Ahn, K.; Gallucci, J. C.; Linden, S.-M. *J. Am. Chem. Soc.* **1984**, *106*, 7980–7981.

(3) The absorbance divided by the total concentration of phenol in all forms.

0022-3263/86/1951-0577$01.50/0 © 1986 American Chemical Society

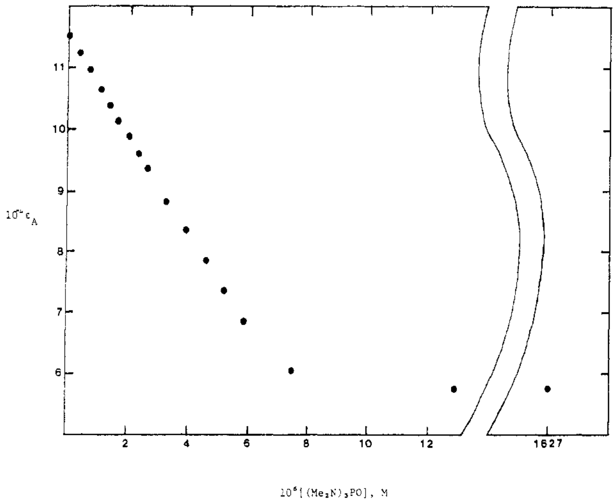


Figure 1. Plot of apparent extinction coefficient of about 6.8 × 10^{-6} M 1,8-biphenylenediol in cyclohexane at 23.5 °C vs. concentration of added *N,N,N′,N′,N″,N″*-hexamethylphosphoramide.

assumption that only a one-to-one adduct is formed gives eq 1, in which ϵ_A is expressed in terms of the total concentrations of base, B, and phenol, P, the extinction coefficients of the free phenol, ϵ_P, and of its complex, ϵ_{PB}, and the equilibrium constant, K, for complex formation.

$$\epsilon_A = \epsilon_P + (\epsilon_{PB} - \epsilon_P)\,(1 + 1/(KP) + B/P - ((1 + 1/(KP) + B/P)^2 - 4B/P)^{1/2})/2 \quad (1)$$

All the points in Figure 1 correspond to complex formation that is so nearly complete that no very reliable equilibrium constant can be calculated from the data. A least-squares treatment[4] using eq 1 gave a K value of 1.25 × 10^7 M^{-1}, but this value is highly dependent on the assumption that we know accurately the relative concentrations of diol and amide, each of which was present in concentrations below 10^{-5} M for most of the points taken. (The solubility of the 1,8-diol in cyclohexane is about 1 × 10^{-5} M at 23 °C). Special precautions had to be taken to avoid complications arising from adsorption of the diol on the walls of the spectrophotometric cell. If we assume that the concentration of diol relative to that of amide is larger than we had originally thought, the calculated value of K increases, and the estimated standard deviations of the calculated absorbances from the observed values decrease, pass through a minimum, and then increase.[5] For the run illustrated in Figure 1 this minimum occurs when the ratio of diol to amide is assumed to be 4% larger than the original value; the estimated standard deviation is 0.001 68, only slightly more than half as large as the original value of 0.003 14; the K value of 2.27 × 10^7 M^{-1} is almost twice the original value. In the other runs that were carried out at 23.5 °C the minimum estimated standard deviation for absorbances is obtained when the diol–amide ratio is taken to be larger by 2%, 6%, and 6% than the original values. We are more confident that the diol–amide complex has a 1:1 (rather than 1.04:1 or 1.06:1) stoichiometry than we are that we know the relative concentrations of the two species with an accuracy greater than 6%. For this reason we have taken the best values of K for complexing of the 1,8-diol with hexamethylphosphoramide to be those obtained with the assumed ratio of diol to amide that gave the smallest standard deviation for the absorbance values.

Table I. Equilibrium Constants for Complexing of 1,8-Biphenylenediol with Various Bases in Cyclohexane

	at 23.5 °C[a]		at 43 °C[a]	
base	K, M^{-1}	$10^{-4}\epsilon_{PB}$[b], M^{-1} cm^{-1}	K, M^{-1}	$10^{-4}\epsilon_{PB}$[b], M^{-1} cm^{-1}
1,3-dioxolane	20	8.86		
	(3)	(0.11)		
EtCN	36	8.87		
	(5)	(0.10)		
1,4-dioxane	73	8.51	39	8.37
	(8)	(0.31)	(1)	(0.06)
Et_2O	134	7.68		
	(6)	(0.07)		
$c\text{-}C_3H_5NH_2$	192	5.76		
	(24)	(0.36)		
tetrahydrofuran	245	7.65		
	(13)	(0.41)		
Me_2NCN	629	6.64		
	(28)	(0.11)		
$HCONMe_2$	12 900	5.92	4610	6.41
	(1200)	(0.04)	(190)	(0.08)
$(Me_2N)_2CO$	58 100	5.57		
	(10 800)	(0.13)		
1,3-dimethyl-2-imidazolidinone	94 900	5.70		
	(11 100)	(0.07)		
Me_2SO	139 000	5.54		
	(21 000)	(0.19)		
tetramethylene sulfoxide	612 000	5.83		
	(48 000)	(0.15)		
$(Me_2N)_3PO$	2.1 × 10^7	5.37	5.3 × 10^6	5.64
	(0.6 × 10^7)	(0.05)	(1.1 × 10^6)	(0.13)

[a] ±0.5 °C. The parenthesized figures are estimated standard deviations. [b] At 261 nm.

Smaller equilibrium constants do not give such a clear picture of the stoichiometry of complex formation, and the standard deviations of the absorbances and calculated equilibrium constants are much less sensitive to assumptions made about the exact ratio of reactants present. Only the equilibrium constants for complexing of 1,8-diol with hexamethylphosphoramide were calculated from the reactant ratios that minimized the standard deviation of the absorbances.

Equilibrium constants for complexing of the 1,8-diol with several other bases at about 23.5 °C and, in some cases, 43 °C were also determined by measurements at 261 nm. The values of K and ϵ_{PB} and the estimated standard deviations[6] obtained are listed in Table I. The value of ϵ_P obtained at 23.5 °C was in the range (1.14 ± 0.02) × 10^5 M^{-1} cm^{-1} in all cases except that of hexamethylphosphoramide (HMPA), where adjusting the reactant ratio caused it to be (1.10 ± 0.02) × 10^5 M^{-1} cm^{-1}. Not surprisingly, the bases with larger equilibrium constants for hydrogen bonding tend to give complexes in which the extinction coefficient has been changed more relative to that of the 1,8-diol. This is seen in Figure 2, in which ϵ_{PB} has been plotted against log K. The correlation is better if only oxygen bases are compared with each other. The line drawn through the points for oxygen bases is arbitrary. The increase in ϵ_{PB} with decreasing log K should not continue indefinitely. It is more reasonable that ϵ_{PB} should approach ϵ_P as K approaches zero and that it should approach the extinction coefficient of the dianion of the diol as K approaches infinity.

---

[4] Hamilton, W. C. "Statistics in Physical Science"; Ronald Press Co.: New York, 1964; Section 5–3.

[5] In the calculations reported here we take the original amide concentration, based on weighing and quantiative dilution, as being correct and adjust the concentration of diol, whose original concentration had been obtained from an extinction coefficient, based on weighing and dilution. Adjusting the amide concentration instead gives essentially the same result.

[6] Estimated standard deviations were taken as $(\sigma_i^2 + \sigma_e^2)^{1/2}$ where σ_i is the average of the standard deviations, calculated as described in ref 4, obtained in the individual runs, and σ_e is the standard deviation of the least-squares values of K or ϵ_{PB} obtained in individual runs from their means.

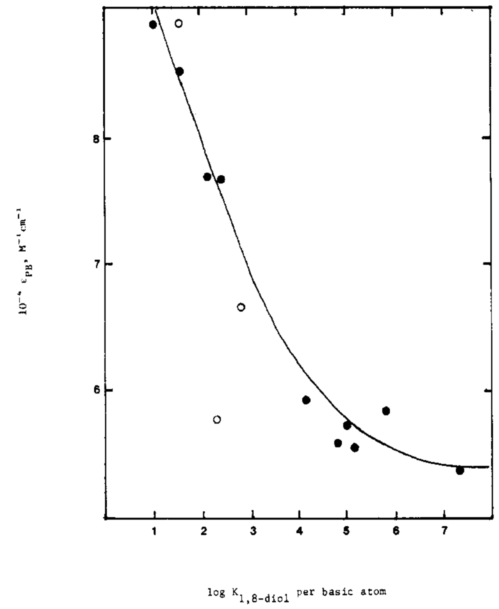


Figure 2. Plot of extinction coefficients of hydrogen bonded complexes of 1,8-biphenylenediol with various bases vs. log K for complex formation: (●) oxygen bases; (○) nitrogen bases.

Table II. Equilibrium Constants for Complexing of *m*-Nitrophenol with Various Bases in Cyclohexane

	at 23.5 °C^a		at 43 °C^a	
base	K, M^{-1}	$10^{-3}\epsilon_{PB}$,^{b,c} M^{-1} cm^{-1}	K, M^{-1}	$10^{-3}\epsilon_{PB}$,^{b,c} M^{-1} cm^{-1}
1,3-dioxolane	12	4.09		
	(2)	(0.04)		
EtCN	42	3.67		
	(3)	(0.06)		
1,4-dioxane	38	4.13	25	4.08
	(2)	(0.02)	(1)	(0.04)
Et_2O	64	3.94		
	(5)	(0.14)		
c-$C_3H_5NH_2$	220	3.42		
	(52)	(0.02)		
tetrahydrofuran	102	3.73		
	(5)	(0.16)		
Me_2NCN	443	3.48		
	(19)	(0.02)		
$HCONMe_2$	1420	3.36	695	3.54
	(120)	(0.02)	(32)	(0.02)
$(Me_2N)_2CO$	4100	3.24		
	(70)	(0.13)		
1,3-dimethyl-2-imidazolidinone	5320	3.22		
	(170)	(0.15)		
Me_2SO	6380	3.33		
	(550)	(0.04)		
tetramethylene sulfoxide	20 500	3.22		
	(900)	(0.02)		
$(Me_2N)_3PO$	406 000	2.94	156 000	3.02
	(61 000)	(0.05)	(21 000)	(0.03)

[a] ±0.5 °C. The parenthesized figures are estimated standard deviations. [b] At 252 nm. [c] $\epsilon_P = (5.35 \pm 0.07) \times 10^3$ M^{-1} cm^{-1}.

The absorption maximum for the complex was also correlated, also imperfectly, with the equilibrium constant, ranging from 262.8 nm for dioxolane to 265.9 nm for hexamethylphosphoramide. Correlations of hydrogen bond strengths with characteristics of absorption maxima in electronic spectra have been discussed by a number of other workers.[7-9a]

Table III. Equilibrium Constants for Complexing of Various Bases with *p*-Nitrophenol and *p*-Fluorophenol in Cyclohexane Solution^a

	p-nitrophenol at 23.5 °C^b		*p*-fluorophenol at 25 °C^c
base	K, M^{-1}	$10^{-4}\Delta_{PB}{}^d$, M^{-1} cm^{-1}	K, M^{-1}
1,3-dioxolane	16	0.78	
	(3)	(0.01)	
EtCN	49	0.94	16^{e,f}
	(8)	(0.01)	
1,4-dioxane	46	0.78	7
	(5)	(0.01)	
Et_2O			13^e
c-$C_3H_5NH_2$			51
tetrahydrofuran	124	0.90	27^e
	(3)	(0.01)	
Me_2NCN	650	1.00	
	(65)	(0.01)	
$HCONMe_2$	2240	1.06	200
	(70)	(0.01)	
$(Me_2N)_2CO$	7340	1.10	370^e
	(400)	(0.01)	
Me_2SO	10 300	1.05	360
	(200)	(0.01)	
$(Me_2N)_3PO$			6300

[a] The parenthesized figures are estimated standard deviations. [b] Measured in the present study. [c] Data from ref 10. [d] At 310 nm. $\epsilon_P = (2.81 \pm 0.15) \times 10^3$ M^{-1} cm^{-1}. [e] Estimated from data in carbon tetrachloride as described in the text. [f] Assumed to be the same as for acetonitrile.

In Table II are values of K and ϵ_{PB} for *m*-nitrophenol and the same set of bases. In Table III are values of K for *p*-nitrophenol and *p*-fluorophenol and most of the bases. Most of the values for *p*-fluorophenol are literature values obtained in cyclohexane solution.[10] The remaining values were obtained from values in carbon tetrachloride[10,11] by use of the literature linear free energy relationship between log K values in the two solvents.[10] This gave 37–110% larger values in cyclohexane than in carbon tetrachloride. For both *m*- and *p*-nitrophenol ϵ_{PB} and λ_{max} for PB correlated, albeit imperfectly, with K. The reliability of all the equilibrium constants depends on the assumption of one-to-one stoichiometry. Possible complications arising from two-to-one complex formation by the diol will be considered later.

Comparison of Equilibrium Constants Obtained with Different Phenols. Taft and co-workers have shown that a log–log plot of equilibrium constants for hydrogen bonding of bases to a given meta- or para-substituted phenol against the corresponding constants for a second such phenol gives a straight line.[11] As the first phenol is made increasingly more acidic than the second the slope of the line becomes larger and larger than 1.0. If 1,8-biphenylenediol is hydrogen bonding with only one of its hydroxy groups at a time, its equilibrium constants for hydrogen bonding to a given base should be about the same, per hydroxy group, as the equilibrium constant for hydrogen bonding of the same base to a phenol of about the same acidity as the diol. The diol's pK_a value of 8.01 corresponds to 8.31 on a per hydroxy group basis. The pK_a for *m*-nitrophenol is 8.36.[12] The data in Tables I and II

---

(7) Pimentel, G. C.; McClellan, A. L. "The Hydrogen Bond"; W. H. Freeman: San Francisco, 1960; Section 4.2.

(8) Vinogradov, S. N.; Linnell, R. H. "Hydrogen Bonding"; Van Nostrand Reinhold: New York, 1971; Section 4.8.

(9) Joesten, M. D.; Schaad, L. J. "Hydrogen Bonding"; Marcel Dekker, Inc.: New York, 1974; (a) Section 1, V; (b) Section 5, II, 1, a.

(10) Joris, L.; Mitsky, J.; Taft, R. W. *J. Am. Chem. Soc.* **1972**, *94*, 3438–3442.

(11) Taft, R. W.; Gurka, D.; Joris, L.; Schleyer, P. v. R.; Rakshys, J. W. *J. Am. Chem. Soc.* **1969**, *91*, 4801–4808.

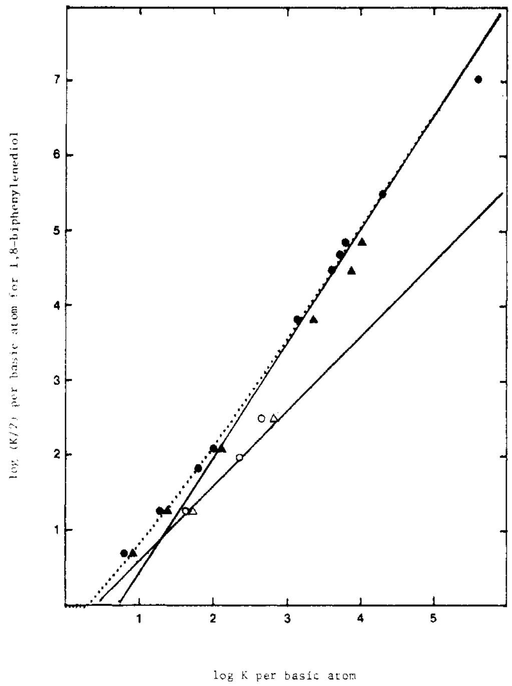


Figure 3. Log–log plots of equilibrium constants per basic atom in cyclohexane at 23.5 °C for hydrogen bonding of 1,8-biphenylenediol to various bases vs. equilibrium constants per basic atom for hydrogen bonding of *m*-nitrophenol and *p*-nitrophenol to the same bases. Circular symbols are for *m*-nitrophenol and triangular symbols for *p*-nitrophenol. Solid symbols are for oxygen bases and open symbols for nitrogen bases.

for oxygen bases show that, for the two weakest bases at 23.5 °C, $K/2$ values for the diol are not quite as large as the corresponding K values for the slightly less acidic *m*-nitrophenol, but the differences are probably smaller than the experimental uncertainties. With the more strongly basic oxygen bases, however, $K/2$ for the diol becomes increasingly larger than K for *m*-nitrophenol and even larger than K for the substantially more acidic *p*-nitrophenol (pK_a 7.15)[12] as shown in Table III. These facts are illustrated in Figure 3, a log–log plot of $K/2$ per basic atom (meaning $K/4$ for dioxolane and dioxane) for diol vs. K per basic atom ($K/2$ for dioxolane and dioxane) for *m*- and *p*-nitrophenol. The slopes of the least-squares best lines (not shown in the figure) through the two sets of points for oxygen bases are 1.37 for *m*-nitrophenol and 1.31 for *p*-nitrophenol. If the diol were forming single hydrogen bonds, the slopes should be about 1.00 and 0.94 (the slope of a plot of the *m*-nitrophenol vs. the *p*-nitrophenol), respectively. If the diol were forming two hydrogen bonds to the bases that were completely independent of each other, the slope of the plot of diol vs. *m*-nitrophenol should be about 2.0. However, hydrogen bonding is a partial proton transfer, and, therefore, the formation of one hydrogen bond to an oxygen atom must reduce the basicity of that oxygen atom. In addition the first hydrogen bond can be formed with the optimum C═O···H—O bond angle for hydrogen bonding. The rigid geometry of the diol probably makes it impossible for both hydrogen bonds to have the optimum bond angle. Finally hydrogen bonding by one hydroxyl group of the 1,8-diol should decrease the acidity, and, hence, decrease the hydrogen bonding ability of the remaining hydroxy group.

It would be expected that nitrogen bases, having only one unshared pair of electrons, would be incapable of forming double hydrogen bonds to the diol or that they would form weaker double hydrogen bonds than they would if their basic atoms had two pairs of unshared electrons. This expectation agrees with the fact that with nitrogen bases $K/2$ for the diol is always smaller than K for *m*- or *p*-nitrophenol. The slopes of the least-squares best lines for statistically corrected plots of the diol data vs. those for *m*- and *p*-nitrophenol, 1.18 and 1.11, respectively, are both larger than 1.0, although they are smaller than those for the corresponding plots of data on oxygen bases. This suggests that the diol is forming double hydrogen bonds to the nitrogen bases (or, at least, to the stronger nitrogen bases), but the lines for nitrogen bases are based on so few data, and these data are sufficiently uncertain, that this is not a very firm conclusion.

Table IV. Equilibrium Constants for Complexing of Bases with 1-Biphenylenol and 8-Methoxy-1-biphenylenol[a]

	1-biphenylenol		8-methoxy-1-biphenylenol	
base	K, M^{-1}	$10^{-4}\epsilon_{PB}{}^{b}$, M^{-1} cm^{-1}	K, M^{-1}	$10^{-4}\epsilon_{PB}{}^{c}$, M^{-1} cm^{-1}
1,4-dioxane	14	8.12	8.6	7.82
	(2)	(0.25)	(1.2)	(0.12)
tetrahydrofuran	32	7.80	20	8.58
	(3)	(0.18)	(3)	(0.07)
$HCONMe_2$	241	7.06	148	9.99
	(22)	(0.10)	(14)	(0.12)
$(Me_2N)_3PO$	25100	6.48	15200	10.30
	(3900)	(0.22)	(500)	(0.06)

[a] Values at 23.5 ± 0.5 °C in cyclohexane solution. The parenthesized figures are estimated standard deviations. [b] At 251 nm. $\epsilon_P = (9.16 \pm 0.09) \times 10^4 M^{-1} cm^{-1}$. [c] At 266 nm. $\epsilon_P = (4.56 \pm 0.13) \times 10^4 M^{-1} cm^{-1}$.

If the diol were a meta- and/or para-substituted phenol, not only should a plot of the data vs. those for *m*-nitrophenol have a slope near 1.00 but also all the $K/2$ values for the diol should be essentially equal to the corresponding K values for *m*-nitrophenol. That is, all the points should fall near a line of slope 1.0 passing through the origin. To check the hypothesis we have already made, that single hydrogen bonding in the diol is sterically hindered by the rather substantial "ortho" substituent (or perhaps by some other factor characteristic of 1-biphenylenol derivatives), we have made measurements with two other derivatives of 1-biphenylenol. We have studied 8-methoxy-1-biphenylenol, in which the ortho substituent is somewhat larger, and 1-biphenylenol, in which the ortho substituent is somewhat smaller than it is with the diol. Equilibrium constants were measured for these 1-biphenylenols complexing with several oxygen bases with the results shown in Table IV. All the resulting values are smaller than those for the same bases complexing with the 1,8-diol or with *m*-nitrophenol, but both of these 1-biphenylenols are less acidic than the 1,8-diol or *m*-nitrophenol. The plots of log K per basic atom for these compounds vs. log K per basic atom for *m*-nitrophenol in Figure 4 show slopes somewhat less than 1.0 for 1-biphenylenol (0.82) and 8-methoxy-1-biphenylenol (0.81). This normal behavior supports the reliability of the individual experimental values. In Figure 5 is a plot of log K for *N*,*N*-dimethylformamide complexing with these 1-biphenylenols and several other phenols vs. the pK_a values for the phenols. The pK_a values for *p*-fluorophenol, 1-biphenylenol, and 8-methoxy-1-biphenylenol are 9.90,[12] 8.64,[13] and 9.15.[13] The least-squares line through the three points for meta- and para-substituted phenols has a slope of −0.39. The

---

[12] Serjeant, E. P.; Dempsey, B. "Ionization Constants of Organic Acids in Aqueous Solution"; Pergamon Press: New York, 1979.

[13] Hine, J.; Hahn, S.; Miles, D. E.; Ahn, K. *J. Org. Chem.* **1985**, *50*, 5092–5096.

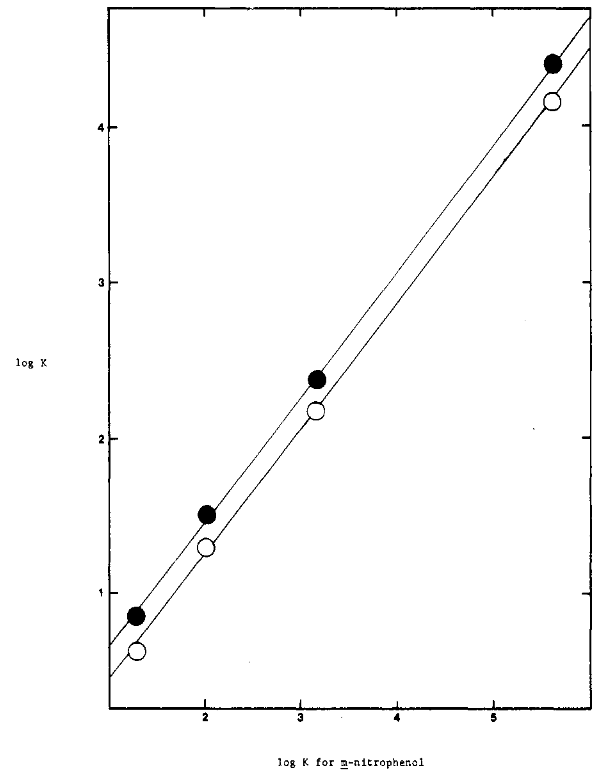


Figure 4. Log–log plots of equilibrium constants per basic atom in cyclohexane at 23.5 °C for hydrogen bonding of 1-biphenylenol (solid circles) and 8-methoxy-1-biphenylenol (open circles) to various bases vs. equilibrium constants per basic atom for hydrogen bonding of *m*-nitrophenol to the same bases.

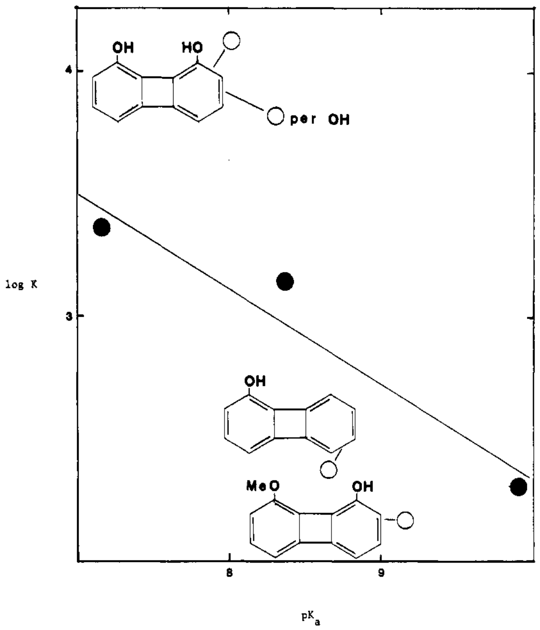


Figure 5. Plots of log K for hydrogen bonding of several phenols to *N,N*-dimethylformamide in cyclohexane at 23.5 °C vs. the pK_a values for the phenols. Solids circles are for meta- and para-substituted phenols. Open circles are used for 1-biphenylenols as indicated.

values of log K for 1-biphenylenol and 8-methoxy-1-biphenylenol lie below this line by 0.48 and 0.50, respectively. An analogous plot for tetrahydrofuran had a slope of −0.25 and negative deviations of 0.30 and 0.37 for 1-biphenylenol and its 8-methoxy derivative, respectively, whereas the slope of the dioxane plot was −0.20, and the negative deviations were 0.28 and 0.39, respectively. If we assume that if the 1,8-diol were singly hydrogen bonding it would give the same type of behavior and ignore hints that the deviations may be increasing with increasing K, we may say that log (K/2) for the diol would deviate by 0.39 ± 0.11 from the value for an meta- or para-substituted phenol of the same acidity. This gives us eq 2 for single hydrogen bonding by the diol. For double hydrogen bonding, the

$$\log (K_{\text{S-1,8}}/2) = \log K_{m\text{-O}_2\text{NPhOH}} - 0.39 \qquad (2)$$

appropriate relationship is eq 3, and,

$$\log (K_{\text{D-1,8}}/2) = \alpha \log K_{m\text{-O}_2\text{NPhOH}} + C \qquad (3)$$

therefore, the total equilibrium constant for hydrogen bonding by the 1,8-diol is expressed in eq 4. All the K

$$K_{\text{T-1,8}}/2 = 10^{-0.39}K_{m\text{-O}_2\text{NPhOH}} + 10^{C}K^{\alpha}{}_{m\text{-O}_2\text{NPhOH}} \qquad (4)$$

values in eq 2–4 are per basic atom. Values of α and C (1.53 and −1.06, respectively) were obtained by fitting all the K values for the diol except the relatively unreliable value for hexamethylphosphoramide to eq 4. From these values the nine $K_{\text{T-1,8}}$ values could be calculated with a standard deviation of 14%. The steeper straight line in Figure 3 is a plot of eq 3 for double hydrogen bonding with these C and α values, and the less steep straight line is a plot of eq 2 for single hydrogen bonding. Equation 4, which represents the sum of these two types of hydrogen bonding, is plotted as a dotted line. The agreement of this curved line with the nine experimental points is considerably better than that of the least-squares straight line (slope 1.40) through the same nine points, where the standard deviation is 26%. This suggests that the stronger oxygen bases are forming double hydrogen bonds almost exclusively, but that a major fraction of the dioxolane and dioxane is forming only a single hydrogen bond to the 1,8-diol. The fact that the line for single hydrogen bonding, which was obtained purely from data on oxygen bases, passes within the experimental uncertainty of the points for the two weaker nitrogen bases supports the assumption that these bases are also singly hydrogen bonding. The fact that the equilibrium constant for complexing of dimethylcyanamide with the 1,8-diol is 74% too large to fall on this line suggests that this nitrogen base may be involved in significant amounts of double hydrogen bonding.

For those bases for which equilibrium constants were measured at 43 °C, values of ΔH and ΔS are listed in Table V along with literature values for *p*-fluorophenol in carbon tetrachloride at 25 °C. The equilibrium constants for double hydrogen bonding might be expected to give more negative entropies than those for single hydrogen bonding because of the extra interference with internal rotations around bonds that should accompany formation of the doubly hydrogen-bonded adduct. Although there seems to be a trend in this direction, the differences in ΔS values are within the experimental uncertainty. Arnett and coworkers, who made calorimetric measurements as well as equilibrium constant determinations to obtain more precise ΔH and ΔS values for *p*-fluorophenol in carbon tetrachloride at 25 °C, found a rather good correlation of $T\Delta S$ with ΔH for all 12 of the ketones, amines, ethers, and esters studied. Increases of ΔH were accompanied by increases in $T\Delta S$. However, this was not so for pyridines, all five of which had ΔS values in the range −14.8 ± 1.4 eu, and the seven sulfoxides, amides, and phosphine oxides, all of which had ΔS values in the range −11.7 ± 2.2 eu.[14]

State of Aggregation of 1,8-Biphenylenediol. Equilibrium constants for association of phenols via hydrogen bonding have been reported, but the determination

---

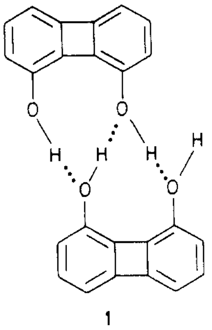
[14] Arnett, E. M.; Joris, L.; Mitchell, E.; Murty, T. S. S. R.; Gorrie, T. M.; Schleyer, P. v. R. *J. Am. Chem. Soc.* **1970**, *92*, 2365–2377.

Table V. Thermodynamics of Hydrogen Bonding of Several Phenols to Various Bases

	1,8-biphenylenediol[a]			m-nitrophenol[a]			p-fluorophenol[b]		
base	$-\Delta G^c$, kcal mol^{-1}	$-\Delta H$, kcal mol^{-1}	$-\Delta S$, eu	$-\Delta G^c$, kcal mol^{-1}	$-\Delta H$, kcal mol^{-1}	$-\Delta S$, eu	$-\Delta G$, kcal mol^{-1}	$-\Delta H$, kcal mol^{-1}	$-\Delta S$, eu
1,4-dioxane	2.53	5.9	11	2.14	4.0	6	1.36[d]	5.10[e]	12.5[f]
	(0.06)	(1.1)	(4)	(0.03)	(0.6)	(2)		(0.11)	
$HCONMe_2$	5.58	9.8	14	4.28	6.8	9	2.81	6.6	12.7
	(0.05)	(1.0)	(3)	(0.05)	(0.9)	(3)	(0.01)	(0.1)	(0.3)
$(Me_2N)_3PO$	9.94	13.3	11	7.61	9.1	5	4.85	8.0	10.6
	(0.15)	(3.5)	(11)	(0.08)	(1.8)	(6)	(0.02)	(0.1)	(1.3)

[a] From data at 23.5 °C and 43 °C and the Van't Hoff equation. In cyclohexane solution. Parenthesized figures are estimated standard deviations. [b] Data from ref 14 obtained by combining equilibrium measurements and calorimetric measurements in carbon tetrachloride solution at 25 °C unless otherwise indicated. [c] At 23.5 °C. [d] From the equilibrium constant in ref 10. [e] Value obtained by using the "pure base" method (in ref 14). [f] Calculated from ΔG and ΔH.

of such constants is complicated by the possibility that significant amounts of dimer, trimer, tetramer, etc., may be present at the same time. Nevertheless, the data avilable on monohydroxylic phenols[9b] show that not more than 1% of the phenol should be associated via hydrogen bonding between the hydroxy groups in a poor hydrogen-bonding solvent at a total phenol concentration below 0.001 M. None of the equilibrium constants we measured used a phenol at this large a concentration. Since the nitro group is only a weakly basic group in hydrogen bonding,[11] no extra large association would be expected from the nitrophenols we studied. All three of the phenols we used gave good Beer-Lambert plots up to the highest concentrations we used. Nevertheless, for the 1,8-diol a formula for a dimer may be written that contains three hydrogen bonds as seen in structure 1. We initially took decreases in the apparent extinction coefficient of the 1,8-diol that were found to accompany decreases in diol concentration below 10^{-6} M as indicating a large equilibrium constant for dimerization.[1b] However, it was later found that these findings were in error because of adsorption of the diol on the walls of containers, including spectrophotometric cells.



1

A triply hydrogen-bonded structure may also be written for the dimer of 2,7-dimethyl-1,8-biphenylenediol, whose molecular weight has been found, by vapor pressure osmometry, to be 204 (calcd 212) at a concentration of 0.0025 M in toluene solution at 56 °C.[15] The unsubstituted diol is not soluble enough in any solvent to which it does not hydrogen bond fairly strongly to make useful molecular weight determinations possible. A doubly hydrogen-bonded structure rather similar to 1 may be written for the dimer of 8-methoxy-1-biphenylenol. If 1 is highly stable, 8-methoxy-1-biphenylenol might be expected to dimerize to a detectable extent at the fairly high concentrations at which it can be studied in non-hydrogen-bonding solvents. However, ultraviolet measurements in cyclohexane solution over the concentration range 1.7×10^{-3} to 1.7×10^{-6} M gave no trend in apparent extinction coefficients. The ^1H NMR chemical shift of the hydroxylic proton in carbon tetrachloride was 4.27 at 0.0160 M and 4.60 at 0.201 M. The chemical shift of the hydroxylic proton of phenol changed from 4.36 to 5.02 in carbon tetrachloride over about the same concentration change. Apparently 8-methoxy-1-biphenylenol has no more tendency to dimerize than phenol does. This tendency has already been stated to be negligible at the concentrations at which we studied phenols spectrophotometrically.

The strongest evidence that the 1,8-diol is at least largely monomeric in cyclohexane solution was obtained by varying the concentration of diol used in determining the equilibrium constant for complexing with a base. The equation we have used to calculate equilibrium constants is based on the assumption that the diol is monomeric and that the complex has a one-to-one stoichiometry. If the diol were largely dimeric, we would get systematic deviations of our calculated from observed apparent extinction coefficient as the amount of added base increased and the concentration of remaining diol decreased. We did not observe such systematic deviations, but in order to be sure we determined the equilibrium constant for complexing with tetrahydrofuran at an initial diol concentration of 3.5×10^{-7} M (instead of the usual $\sim 7 \times 10^{-6}$ M) in a 10-cm cell (instead of the usual 1-cm cell). The value obtained was about 5% larger than the average of the values obtained in a 1-cm cell. If the diol were almost all dimer, it should have been about $20^{1/2}$ times as large at a concentration 20-fold smaller.

Possible Complications Arising from Complexing of Diol with Two Molecules of Base. To examine the possibility of the formation of a complex between one molecule of diol and two molecules of base (a PB_2) we made a plot of ϵ_A vs. log [base] for each of our bases. For none was there a clear indication of the kind of break that would come from two separable consecutive equilibria. We would not expect to see such a break if ϵ_{PB} were approximately equal to ϵ_P or to ϵ_{PB}, or if the ratio of the two equilibrium constants were less than about 50, or if the amount of PB_2 formed was small. If any of these three alternatives is correct we would not be able to obtain two reliable ϵ values and two reliable K values from the data. Hence, we did not try.

Plots of ϵ_A vs. wavelength showed fairly good isosbestic points (near 262 and 257 nm) for the six strongest bases for the entire range of base concentrations studied. For the other bases there was a satisfactory isosbestic point near 262 nm over the range of base concentrations used in calculating K. There were deviations that are probably significant at 262 nm at higher base concentrations and deviations near 257 nm for some of the higher base concentrations used in calculating K. Isosbestic points must exist, if the spectra cross each other, in the concentration range where PB_2 formation is negligible.

---

[15] Molecular weight determined by Dr. C.-S. Wu of Ashland Chemical Company, to whom we express our gratitude.

In addition to the ϵ_A values at 261 nm, near the 260.5-nm absorption maximum for the 1,8-diol, where ϵ_A always decreased when base was added, ϵ_A was also measured at 265 nm, nearer the absorption maxima for the complexes, where it was increased (initially, at least) when base was added. However, there is evidence for significantly interfering absorbance by PB_2 at this wavelength. Especially for the weaker bases, the largest amounts of added base caused ϵ_A to increase more at 265 nm than would be expected from the smaller decreases in ϵ_A at 261 nm. Consequently, the standard deviation of the absorbance values calculated by using the least-squares K value from the observed absorbance values were as large (for tetramethylene sulfoxide) or larger when absorbances at 265 nm were used as when absorbances at 261 nm were. Except for hexamethylphosphoramide, for which the calculations were not made, they averaged 75% larger. These standard deviations of the absorbances were less than 0.004 for the eight stronger oxygen bases when 261-nm data were used. They were larger than 0.004 for the two weaker oxygen bases and the three nitrogen bases when 261-nm data were used and for nine of the ten weakest bases when 265-nm data were used. The K values obtained from 265-nm data were smaller than the corresponding 261-nm values by 30–46% for the five weakest bases and tetramethylene sulfoxide and by 13–25% for the other six bases. We believe that the 261-nm absorbances are less contaminated by PB_2 absorbance and have therefore, used these values. If we instead used the K values obtained at 265 nm, however, none of our qualitative conclusions would change.

The equilibrium constant for single hydrogen bonding, which we will write as K_S to be more general than we were in eq 2, may be expressed as shown in eq 5, in which P_e

$$K_S = \frac{C_{Se}}{P_e B_e} \tag{5}$$

and B_e are the equilibrium concentrations of phenol and base, and C_{Se} is the equilibrium concentration of that part of the total one-to-one complex that is singly hydrogen bonded. The equilibrium constant K_2 for addition of a second molecule of base to the singly hydrogen-bonded one-to-one complex to give the two-to-one complex is defined by eq 6, in which C_{2e} is the equilibrium concentration of the two-to-one complex. The total equilibrium constant

$$K_2 = \frac{C_{2e}}{B_e C_{Se}} \tag{6}$$

for one-to-one complex formation, K_T, bears the same relationship to K_S that C_{Te}, the total concentration of one-to-one complex, bears to C_{Se} (eq 7). From this, eq 8

$$\frac{K_T}{K_S} = \frac{C_{Te}}{C_{Se}} \tag{7}$$

for K_E, the equilibrium constant for complexing of the actual one-to-one complex with a second molecule of base,

$$K_E = \frac{C_{2e}}{B_e C_{Te}} = \frac{K_S K_2}{K_T} \tag{8}$$

may be derived. The K in eq 1 we intend to calculate is K_T, for one-to-one complex formation. What can be said about interference in this calculation from formation of two-to-one complex via K_E? Let us first consider the relative magnitude of K_S and K_2, the equilibrium constants for complexing by a base, via a single hydrogen bond, with the first and second hydroxy groups of a symmetrical diol. Even if the presence of the first base does not affect the hydrogen bonding acidity of the remaining hydroxy group, K_E will equal $K_T/4$ because of statistical effects.

However, K_E would be expected to be reduced further, both by steric and by polar effects. Many observations, such as Figure 5, show that decreasing the acidity of a phenol tends to decrease its hydrogen bonding acidity. The partial deprotonation of the diol, resulting from hydrogen bonding to one hydroxy group, must decrease the acidity of the remaining hydroxy group.

We have estimated the polar effect on K_E as follows. We have assumed that the slope α of a plot of log K for hydrogen bonding of a given base to various phenols vs. log K_a for the phenols is linearly related to log K for hydrogen bonding of these bases to *m*-nitrophenol. From the α values we determined experimentally (cf. Figure 5) for three bases, values ranging from 0.14 for dioxolane to 0.64 for hexamethylphosphoramide were obtained. We take α to be proportional to the fraction of proton transfer in the hydrogen bonds. Since complete deprotonation of the 1,8-diol decreases the acidity of the remaining hydroxy group by 2.39 pK_a units, we assume that deprotonation by the fraction α should decrease it by an amount proportionate to 2.39α. Actually α would be much smaller if it were based on pK_a values in cyclohexane rather than in water, because polar substituent effects will be larger in cyclohexane, but correspondingly the 2.39 change in pK_a would be much larger if it were in cyclohexane. Therefore, the effects on the product 2.39α should tend to cancel. If the 2.39α change in pK_a is then multiplied by the factor α which measures the sensitivity of equilibrium constants for hydrogen bonding to changes in pK_a, we get an estimate of $-2.39\alpha^2$ for the polar effect on log K_E.

If K_E is as small as we have estimated, there will be less than 1% transformation of PB to PB_2 at the highest concentration of base used in calculating our K values for the five strongest oxygen bases and only 4% PB_2 for dimethylformamide. The maximum extent of estimated PB_2 formation in the points used to calculate K for the four weakest oxygen bases ranged from 25% to 39% (for dioxolane); for the nitrogen bases it was 37% for dimethylcyanamide, 78% for cyclopropylamine, and 71% for propionitrile.

To test for uncertainties owing to PB_2 formation, we omitted some of the absorbances obtained at the highest base concentrations, where PB_2 formation should be greatest from our least-squares calculation of K values. For a run on each base except hexamethylphosphoramide the three highest base concentrations were omitted, so that the highest base concentration left was smaller by 4- to 60-fold than it had been. There were increases in the K value obtained by 52% for propionitrile, 43% for dioxolane, 22% for dioxane, and 14% for dimethylcyanamide. Some of the other values increased and some decreased but none changed by more than 7%. We take this as evidence that PB_2 formation does not significantly affect the K values obtained, except for the three weakest bases. Since some of these bases whose calculated K values changed by no more than 7% had been used at concentrations where the estimated extent of PB_2 formation was as high as 71%, we believe that K_E values are smaller than we have estimated or that ϵ_{PB_2} is approximately equal to ϵ_{PB}.

Nevertheless, to lessen the effect of PB_2 formation on our calculated K values, we omitted data obtained at some of the higher base concentrations by using the values of K_E we have estimated and the following rules: (1) If the estimated extent of PB_2 formation is always less than 1%, all the data are used. (2) If there is a point with more than

1% PB_2 formation and if dropping this point still leaves a point with 95% or more conversion of P to PB, the point is dropped. (3) If there is a point with more than 2% PB_2 formation and if dropping this point still leaves a point with 90% or more conversion of P to PB, the point is dropped.

Conclusions

Whatever the interpretation of our results on the weaker bases, there is no experimental evidence for any PB_2 formation with the stronger bases and diol. Hence, the relatively large values for the equilibrium constant, the relatively steep slope of the plot of log K for diol vs. log K for *m*-nitrophenol, and the direct determination of the stoichiometry of the complex with hexamethylphosphoramide give very strong evidence that *the stronger oxygen bases form doubly hydrogen-bonded complexes with 1,8-biphenylenediol* almost exclusively.

Experimental Section

Chemicals and Instruments. 1,8-Biphenylenediol was prepared by modification of the method of McOmie et al.,[16,17] and 1-biphenylenol was prepared by a literature method.[18,19] The synthesis of 8-methoxy-1-biphenylenol and 2,7-dimethyl-1,8-biphenylenediol are being reported elsewhere.[13] Spectrophotometric grade cyclohexane (Aldrich) was used directly, or reagent cyclohexane (Fisher) was purified with use of sulfuric acid, then a silica gel column and distillation over calcium hydride. The other phenols and the bases used were ordinarily the best grade commercially available and all were purified further before use.

A Varian DMS 100 UV–vis spectrophotometer equipped with an Epson RX-80 printer or a Cary 1605 spectrophotometer was used.

Determination of Equilibrium Constants. Solubility of the 1,8-diol in cyclohexane was so low and solution was so slow that occasionally heating (not above 50 °C) under argon and shaking was used for 2–3 days to obtain homogeneous solutions. In a typical run the baseline was established by using 3.50 mL of cyclohexane in both the sample cell and the reference cell, and then the contents of the sample cell were replaced by 3.50 mL of 6.765×10^{-6} M 1,8-diol in cyclohexane. After the spectrum had been run from 360 to 240 nm and range 257–267 nm expanded by using the zoom mode of the DMS 100 spectrophotometer, 1 μL of a solution of 5 μL of hexamethylphosphoramide in 25 mL of cyclohexane was added to each of the two cells. When the range 257–267 nm had been scanned again, another 1 μL of hexamethylphosphoramide solution was added. After 40 μL of the solution had been added, in 15 gradually increasing amounts, 1 μL of pure hexamethylphosphoramide was added to each cell, which had been kept at 23.5 ± 0.5 °C throughout the process.

In the runs at 43 °C the cells were cooled for 3 min in a refrigerator before adding each new sample of base solution to minimize volatilization of the cyclohexane solvent. In calculating concentrations at 43 °C allowance was made for thermal expansion of the solutions.

When a solution of about 6×10^{-6} M 1,8-biphenylenediol in cyclohexane was placed in a 1-cm quartz cell, the absorbance dropped by about 2.4% from the value obtained after about 3 s to the constant value reached after about 15 min. Several types of experiments showed that this decrease came from reversible adsorption on the cell walls. Calculations were made, with use of a number of different assumptions as to exactly how this adsorbed diol would go back into solution as base was added, to learn how this would affect the values we have calculated for the equilibrium constants. The K values we have used were estimated to be either too high or too low, by as much as 10%, depending on the assumption used.

In more dilute solutions of diol, especially those used in 10-cm cells, a larger fraction of the diol was adsorbed on the cell walls. Therefore, we devised a method for learning how much diol was actually in solution. A 1-cm cell containing 3.5 mL of diol solution in cyclohexane was allowed to stand until its absorbance was constant. A 3-mL pipet was used to transfer 3 mL of this solution to a 1-cm cell that already contained 1 mL of 3.6 M tetrahydrofuran in cyclohexane, and the mixed solution was drawn up into the pipet several times to rinse it. It is known that only a negligible amount of diol is absorbed from a solution that is 0.9 M in tetrahydrofuran. Then the concentration of diol in the second cell was calculated from the extinction coefficient of diol at 264 nm in 0.9 M tetrahydrofuran in cyclohexane. In this way we were able to use diol concentrations as low as 1.31×10^{-8} M (where the uncertainty in the small absorbance values obtained had reached about 10%).

Registry No. THF, 109-99-9; EtCN, 107-12-0; Et_2O, 60-29-7; c-$C_3H_5NH_2$, 765-30-0; Me_2NCN, 1467-79-4; $HC(O)NMe_2$, 68-12-2; $(Me_2N)_2CO$, 632-22-4; Me_2SO, 67-68-5; $(Me_2N)_3PO$, 680-31-9; 1,8-biphenylenediol, 18798-64-6; *m*-nitrophenol, 554-84-7; *p*-nitrophenol, 100-02-7; 1-biphenylenol, 1078-07-5; 8-methoxy-1-biphenylenol, 98945-48-3; 1,3-dioxolane, 646-06-0; 1,4-dioxane, 123-91-1; 1,3-dimethyl-2-imidazolidinone, 80-73-9; tetramethylene sulfoxide, 1600-44-8.

---

[16] Blatchly, J. M.; Garner, D. V.; McOmie, J. F. W.; Watts, M. L. *J. Chem. Soc. C* **1968**, 1545–1549.

[17] Baker, W.; Barton, J. W.; McOmie, J. F. W. *J. Chem. Soc.* **1958**, 2658–2665.

[18] Boulton, A. J.; Chadwick, J. B.; Harrison, C. R.; McOmie, J. F. W. *J. Chem. Soc. C* **1968**, 328–330.

[19] Prepared by Dr. Shwn-Meei Linden of The Ohio State University, to whom we express our thanks.