A Search for the Addition of Monodentate Nucleophiles to Boric and Benzeneboronic Acids in Water[1]

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A spectrophotometric search for the addition of a variety of monodentate oxygen, nitrogen, sulfur, and carbon nucleophiles to boric and benzeneboronic acids in aqueous solution was made. None of the nucleophiles appeared to react significantly with boric acid. Among the nucleophiles which were tested with benzeneboronic acid, only 1-methylimidazole, imidazole, hydrazine, and probably 2,2,2-trifluoroethoxide ion reacted. The equilibrium constant for the addition of trifluoroethoxide ion to benzeneboronic acid was estimated to be 3500 M^{-1}, at 25° and 1 M ionic strength. The equilibrium constants for the addition of 1-methylimidazole and imidazole were determined spectrophotometrically and, in one case, potentiometrically and have values of 12.2 and 10.7 M^{-1}, respectively, at 25° and 1 M ionic strength. Apparent acid dissociation constants for boric and benzeneboronic acids in the presence of some salts and organic solutes are reported.

Boric acid and boronic acids exist in water as trigonal compounds which ionize by the addition of hydroxide ion to form tetrahedral compounds (eq 1, XR′

$$\mathrm{RB(OH)_2} + \mathrm{XR'} \rightleftharpoons \mathrm{RB(OH)_2XR'} \qquad (1)$$

= OH^-).[2,3] The addition of monodentate nucleophiles other than hydroxide ion to boric and boronic acids in water has received little attention. Edwards has described the reaction of hydroperoxide anion with boric acid (eq 1, XR′ = OOH^-) and determined an approximate equilibrium constant for the reaction.[4] Tanner and Bruice have reported that phenols do not form esters with aqueous boric acid in appreciable amounts; however, they do not give the experimental data upon which this conclusion is founded.[5] In this paper we describe a search for the addition of a variety of monodentate ligands to boric acid and benzeneboronic acid in water. Benzeneboronic acid was found to react with 1-methylimidazole, imidazole, hydrazine, and probably 2,2,2-trifluoroethoxide anion. Except for the reaction with hydrazine, the equilibrium constants for these reactions have been determined. In other cases which we investigated no significant reaction was detected, and therefore we report only upper limits for the equilibrium constants.

Experimental Section

Materials.—Reagent grade boric acid and potassium tetraborate were used without further purification. Benzeneboronic acid from Aldrich Chemical Co. was recrystallized twice from water; it gave the theoretical equivalent weight upon titration with sodium hydroxide in the presence of 0.5 M mannitol.[6] The nucleophiles which were used, with the exceptions of *p*-nitrobenzenethiol, phenol, aniline, and nitromethane, were either recrystallized or redistilled.

Methods.—Spectra were recorded with a Unicam SP800A recording spectrophotometer; individual absorbance readings were taken with either the Unicam or a Gilford 240 spectrophotometer. Teflon-stoppered, 1-cm cuvettes were generally used; in a few cases 5 mm or 1 mm cuvettes were used. pH values were measured with a Radiometer 25 SE pH meter. The meter was standardized just before use with standard buffers at two pH values which encompassed the range of pH measurements. All solutions which were used in the potentiometric determinations of the association constants (see Table II) were prepared with degassed and deionized water saturated with nitrogen and were kept under argon during the measurement of pH. Also, the buffers of benzeneboronic acid which were used for this purpose were made from the acid with carbonate-free potassium hydroxide.[7]

All experiments were carried out with reaction mixtures at 25.0 ± 0.1° and at ionic strength 1.0 M, adjusted with KCl.

Results and Discussion

pK Values of Boric and Benzeneboronic Acids.—Table I presents our values and ones from the literature for the apparent pK's of boric and benzeneboronic acids. The values increase upon the addition of solutes

TABLE I

APPARENT pK VALUES OF BORIC AND BENZENEBORONIC ACIDS, AT 25° AND 1 M IONIC STRENGTH

Boron compd	M	Solvent	pK_{BH}
Boric acid	0.004	Water	9.02[a]
		Water	9.18[b]
	0.004	0.30 M *p*-Dioxane	9.05[a]
	0.004	0.35 M *p*-Dioxane	9.08[a]
	0.004	0.40 M *p*-Dioxane	9.08[a]
	0.004	0.53 M *p*-Dioxane	9.12[a]
	0.004	0.10 M Furan	9.06[a]
		4.3 M Ethanol	9.87[c]
Benzeneboronic acid	0.004	Water	8.68[d]
		Water	8.86[b]
		4.3 M Ethanol	9.70[c]

[a] From potentiometric titrations with 1 N NaOH. The pK was obtained by plotting the data according to the equation pH = pK + log [B]/[HB], where [B] is the concentration of the basic species and [HB] is the concentration of the acid species. In each case a mixture which was identical except for the absence of the boron compound was also titrated with 1 N NaOH. The values of [B] were obtained by subtracting the milliliters of base required to achieve a pH in this control from the milliliters of base required to achieve the same pH in the titration mixture. The values of [HB] were calculated by subtracting the values of [B] at each pH from the end point value of [B]. The plots of pH *vs.* log [B]/[HB] were linear with slopes of 0.93–1.1. [b] In water, at 25° and 0.04 M ionic strength or less, from ref 6. [c] In 4.3 M ethanol, at 25° and 0.04 M ionic strength or less, from reference given in footnote *b*. [d] From a spectrophotometric titration at 267 nm, where the molar extinction coefficients of the acidic and basic species are 435 and 125 M^{-1} cm^{-1}, respectively. For the procedure, see V. Gold, "pH Measurements," Methuen, London, 1956, Chapter VIII. The solutions were buffered at each pH by the benzeneboronic acid itself.

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(1) (a) Supported by Grants GB12848 and GB29205 from the National Science Foundation, a fellowship to K. A. K. from the National Institutes of Health training grant to the Department of Biochemistry and Molecular Biology, and a fellowship to R. C. J. from the National Science Foundation. (b) Address inquiries to G. E. L.

(2) R. P. Bell, J. O. Edwards, and R. B. Jones in "The Chemistry of Boron and Its Compounds," E. L. Muetterties, Ed., Wiley, New York, N. Y., 1967, Chapter 4.

(3) J. P. Lorand and J. O. Edwards, *J. Org. Chem.*, **24**, 769 (1959).

(4) J. O. Edwards, *J. Amer. Chem. Soc.*, **75**, 6154 (1953).

(5) D. W. Tanner and T. C. Bruice, *ibid.*, **89**, 6954 (1967).

(6) G. E. K. Branch, D. L. Yabroff, and B. Bettman, *ibid.*, **56**, 937 (1934).

(7) C. W. Davies and G. H. Nancollas, *Nature (London)*, **165**, 237 (1950).

less polar than water. This increase can be explained by the fact that the ionization reaction, $RB(OH)_2 \rightleftharpoons RB(OH)_3^- + H_3O^+$, is one in which a neutral species yields two ions. An additional complication in the case of ethanol is that ethanol can react with the acids to form an ester (eq 2, $R'XH = CH_3CH_2OH$). Because

$$RB(OH)_2 + R'XH \rightleftharpoons RB(OH)XR' + H_2O \qquad (2)$$

the pK values of ethanol and water are about the same,[8] it seems likely that the equilibrium constant for this reaction *as written* is close to one. The qualitative observation that esters of boric acid hydrolyze in water supports this view.[9] Also, the analogous equilibrium constant for the esterification of acetic acid has a value of 3.3 at 25°.[10] The fraction as ester in 4.3 M ethanol therefore probably lies in the 0 to 0.5 range. If the ester has a much higher pK than the acid, its formation alone could account for the rise in pK in ethanol. However, the substitution of OC_2H_5 for OH probably does not alter the pK value by more than 0.3 units, so that the increase in the apparent pK values which is caused by ethanol is probably largely a solvent effect.

Reaction of Benzeneboronic Acid with Trifluoroethanol.—Solutions of pH 8.60 that contained 0.005 M benzeneboronic acid at 1 M ionic strength in water, in 1 M ethanol, and in 1 M 2,2,2-trifluoroethanol were prepared by determining how much NaOH was necessary to adjust trial solutions of the acid to this pH. The absorbances of these solutions at 267 nm, which is the wavelength of maximum absorbance for the acid itself, against blanks which lacked the benzeneboronic acid were 1.465, 1.626, and 1.359, respectively. The ultraviolet spectra were identical in shape. A similar experiment in which the spectra were obtained in water, 2 M ethanol, and 2 M trifluoroethanol, gave a qualitatively similar result and showed that the absorbances did not change over a period of 18 hr. On the other hand, the spectrum of benzeneboronic acid was not altered by 1 M ethanol or trifluoroethanol at pH 2 and 1 M ionic strength. For the reasons discussed above, the most likely explanation of the result with ethanol is that the ethanol increased the apparent pK of benzeneboronic acid by a solvent effect.[11] The molar absorbancies of the benzeneboronic acid and anion at 267 nm in water (ϵ_{HB} and ϵ_B) were determined to be 435 and 125 M^{-1} cm^{-1}, respectively, and therefore the concentrations of acid and base present in the 1 M ethanol were 0.00322 and 0.00178 M and the apparent pK is 8.86.

The result with trifluoroethanol cannot be interpreted unambiguously, since both the solvent effect and the tendency of trifluoroethanol to react with benzeneboronic acid contribute to the result. In order to estimate a value for the equilibrium constant for the addition of trifluoroethoxide anion to benzeneboronic acid (eq 1, $R'X = CF_3CH_2O^-$), we will assume that the solvent effect is the same as that with ethanol and that the equilibrium constant for the esterification reaction (eq 2, $R'XH = CF_3CH_2OH$) is close to unity. This latter assumption means that the fraction as the ester $CF_3CH_2OBC_6H_5(OH)$ is less than about 5%. According to these assumptions, the smaller absorbance in the presence of 1 M trifluoroethanol than in the presence of 1 M ethanol is due to the formation of the species $CF_3CH_2OB^-C_6H_5(OH)_2$, which would be expected to have a molar absorbancy similar to that of benzeneboronate anion. The approximate concentration of this species is given by eq 3, where A_E and A_{TFE} are the absorbances

$$A_E - A_{TFE} = \left[\epsilon_{BH}\frac{a_H}{a_H + K_{BH}} + \epsilon_B\frac{K_{BH}}{a_H + K_{BH}} - \epsilon_B\right] \times [RB(OH)_2XR'] \qquad (3)$$

of the reaction mixtures with ethanol and trifluoroethanol, a_H is the activity of the hydrogen ion, and K_{BH} is the acid dissociation constant of benzeneboronic acid in 1 M ethanol. Using this concentration, we can calculate an approximate equilibrium constant for the formation of the trifluoroethoxide adduct, according to eq 4, where the subscript T signifies the total concentration of all species and $K_{R'XH}$ is the acid dissociation constant of trifluoroethanol (4×10^{-13} at 25°[8]). The estimated value for K_N is 3500 M^{-1}.

$$K_N = \frac{[RB(OH)_2XR']}{[RB(OH)_2][XR']} = \frac{[RB(OH)_2XR']}{([RB(OH)_2]_T - [RB(OH)_2XR']) \times (a_H/K_{BH} + a_H)[R'XH]_T(K_{R'XH}/K_{R'XH} + a_H)} \qquad (4)$$

The equilibrium constant for the corresponding reaction with hydroxide ion (eq 1, $R'X = HO^-$) is equal to K_{BH}/K_W, where K_W is the ion product of water; and therefore it has a value of about 1.4×10^5 M^{-1}. Thus, it appears that the difference in basicity of 3.3 log units between hydroxide ion (the pK of water is 15.7) and trifluoroethoxide ion is associated with a difference in log K_N of only 1.6 log units. This greater sensitivity of basicity to electronic effects means that the proton is more electron demanding than the $RB(OH)_2$ group. This conclusion is the one that is expected from the fact that the boron adduct bears a net negative charge.

Reaction of Benzeneboronic Acid with 1-Methylimidazole and with Imidazole.—The addition of 1-methylimidazole buffers or imidazole buffers to solutions of benzeneboronic acid at the same pH as the buffer caused the spectrum of the benzeneboronic acid to change from predominantly the spectrum of the acidic species to one which resembled that of the benzeneboronate anion. The spectral change was complete within less than 15 sec, and the new spectrum was stable for at least 2 hr.

The apparent equilibrium constants for these reactions were determined spectrophotometrically at several pH values. The absorbances at 267 nm of reaction mixtures which contained a single concentration of benzeneboronic acid in the presence of various concentrations of 1-methylimidazole or imidazole buffer were measured against air and then corrected for the absorbance of the buffer by subtracting the absorbances of blank mixtures which lacked benzeneboronic acid and which were also measured against air. The concentrations of both the acidic and basic species of the buffer were five times or more as large as the concentration of the benzeneboronic acid. Under these

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[8] P. Ballinger and F. A. Long, *J. Amer. Chem. Soc.*, **82**, 795 (1960).

[9] A. Scattergood, W H. Miller, and J. Gammon, Jr., *ibid.*, **67**, 2150 (1945).

[10] W. P. Jencks and M. Gilchrist, *ibid.*, **86**, 4651 (1964).

[11] The effects of ethanol and trifluoroethanol upon the spectrum are probably not due to solvent effects upon the glass electrode, since the pH value of 10^{-3} N HCl in water, 1 M ethanol, and 1 M trifluoroethanol at 25° and 1 M ionic strength were found to be 3.06, 3.08, and 3.09, respectively.

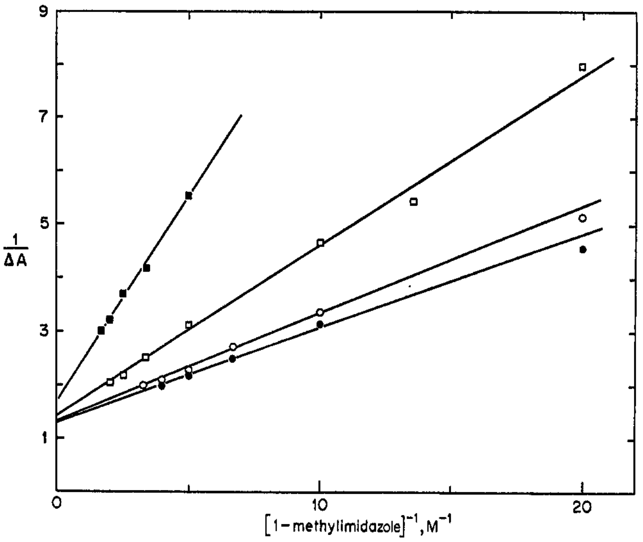


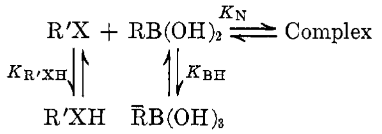
Figure 1.—Spectrophotometric determination of the equilibrium constant for the reaction between 1-methylimidazole and benzeneboronic acid, at 25° and 1 *M* ionic strength. The reciprocal of the difference between the absorbance of benzeneboronic acid alone and the absorbance of benzeneboronic acid in the presence of 1-methylimidazole buffers, at 267 nm, is plotted against the reciprocal of total concentration of 1-methylimidazole buffer. The concentration of benzeneboronic acid was 0.0030 *M* throughout. The 1-methylimidazole buffers contained 20 (■), 40 (□), 60 (○), and 80% (●) free base.

conditions, the equation which describes the formation of a 1:1 complex is eq 5, where ΔA is the difference between the absorbance of the benzeneboronic acid alone at the pH of the buffer and the absorbance of the mixture of benzeneboronic acid and 1-methylimidazole or imidazole buffer; K_{obsd} is the apparent equilibrium constant for complex formation; $[\mathrm{RB(OH)_2}]_{\mathrm{T}}$ is the total concentration of benzeneboronic acid; $\Delta\epsilon$ is the difference between the sum of the molar extinction coefficients of the benzeneboronic acid $[(\epsilon_{\mathrm{BH}}a_{\mathrm{H}} + \epsilon_{\mathrm{B}}K_{\mathrm{BH}})/(K_{\mathrm{BH}} + a_{\mathrm{H}})]$ and the 1-methylimidazole or imidazole buffer ($\epsilon_{\mathrm{R'XH}}$) and the extinction coefficient of the complex (ϵ_{c}); and $[\mathrm{R'XH}]_{\mathrm{T}}$ is the total concentration of the 1-methylimidazole or imidazole buffer. Linear plots of $1/\Delta A$ *vs.* $1/[\mathrm{R'XH}]_{\mathrm{T}}$ were obtained in all cases (see Figure 1 for the results with 1-methylimidazole); and thus the data are adequately explained by the assumption of a 1:1 complex. The values of K_{obsd} increased with the fraction of base in the buffer. This fact suggested that the complex contains the elements of 1-methylimidazole or imidazole base and the acidic species of benzeneboronic acid. The values of the pH-independent equilibrium constant for the formation of the complex from these species (K_{N}) were obtained by use of eq 6, where $K_{\mathrm{R'XH}}$ is the acid dissociation constant of 1-methylimidazolium ion (4.6×10^{-8} *M*) or imidazolium ion (6.0×10^{-8} *M*). The values of K_{N} which were calculated from the various values of K_{obsd} were the same within ±5% of the average value and equal to 12.2 M^{-1} for the complex with 1-methylimidazole and 10.7 M^{-1} for the complex with imidazole.

$$\frac{1}{\Delta A} = \left(\frac{1}{K_{\mathrm{obsd}}[\mathrm{RB(OH)_2}]_{\mathrm{T}}\Delta\epsilon}\right)\left(\frac{1}{[\mathrm{R'XH}]_{\mathrm{T}}}\right) + \frac{1}{[\mathrm{RB(OH)_2}]_{\mathrm{T}}\Delta\epsilon} \qquad (5)$$

$$K_{\mathrm{N}} = K_{\mathrm{obsd}}\left(\frac{K_{\mathrm{BH}} + a_{\mathrm{H}}}{a_{\mathrm{H}}}\right)\left(\frac{K_{\mathrm{R'XH}} + a_{\mathrm{H}}}{K_{\mathrm{R'XH}}}\right) \qquad (6)$$

The equilibrium constant for the reaction of 1-methylimidazole with benzeneboronic acid was also determined by a potentiometric method. Since 1-methylimidazole complexes detectably only with the acidic form of benzeneboronic acid, the addition of 1-methylimidazole base to a buffer of benzeneboronic acid should remove the acidic species and therefore raise the pH to a higher value than the value which would be reached in the absence of complexation. The equilibria which must be considered are given in eq 7,



$$\begin{tikzcd} \mathrm{R'X} \arrow[d, shift left, "K_{\mathrm{R'XH}}"'] & + & \mathrm{RB(OH)_2} \arrow[r, shift left, "K_{\mathrm{N}}"] \arrow[d, shift left, "K_{\mathrm{BH}}"] & \mathrm{Complex} \arrow[l, shift left] \\ \mathrm{R'XH} \arrow[u, shift left] & & \mathrm{\bar{R}B(OH)_3} \arrow[u, shift left] & \end{tikzcd} \qquad (7)$$

where R′XH and R′X refer to the acidic and basic species of the tertiary amine. This scheme yields eq 8,

$$K_{\mathrm{N}} = \frac{K_{\mathrm{BH}}K_{\mathrm{R'XH}}[\mathrm{RB(OH)_2}]_0}{[\mathrm{R'X}]a_{\mathrm{H}}[\mathrm{R\bar{B}(OH)_3}]_0K_{\mathrm{R'XH}} + [\mathrm{R'X}]^2a_{\mathrm{H}}^2} - \frac{K_{\mathrm{BH}}}{[\mathrm{R\bar{B}(OH)_3}]_0K_{\mathrm{R'XH}} + [\mathrm{R'X}]a_{\mathrm{H}}} - \frac{1}{[\mathrm{R'X}]} \qquad (8)$$

in which $[\mathrm{RB(OH)_2}]_0$ and $[\mathrm{RB^{-}(OH)_3}]_0$ are the concentrations of these species in the absence of 1-methylimidazole. Solution of eq 8 for K_{N} requires the value of [R′X], the concentration of the basic species. In our experiments (see Table II) we used total concentrations of 1-methylimidazole which were 15 times those of the benzeneboronic acid. Under these conditions, even if all the boronic acid were complexed with 1-methylimidazole, the concentration of the free base would be equal to 92% or more of the total concentration of 1-methylimidazole. Consequently, we have calculated values of K_{N} from eq 8 by using the total

TABLE II

EQUILIBRIUM CONSTANTS FOR THE REACTION OF TERTIARY AMINES WITH BENZENEBORONIC ACID, FROM POTENTIOMETRIC MEASUREMENTS AT 25° AND 1 *M* IONIC STRENGTH

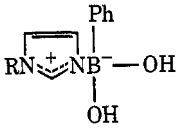
Amine	*M* of $\mathrm{RB(OH)_2}$ buffer	pH of buffer alone	*M* of amine[a]	pH of buffer + amine	K_{N}[b]
1-Methylimidazole	0.01	9.16	0.15	9.72	11.0
	0.02	9.17	0.30	9.89	10.3
	0.01	8.68	0.15	9.37	11.2
	0.02	8.69	0.30	9.51	11.3
	0.02	9.16	0.30	9.91	11.6
	0.023	9.18	0.345	9.96	11.1
	0.01	9.16	0.15	9.72	11.0
Pyridine	0.01	9.15	0.15	9.25	1.6
	0.02	9.16	0.30	9.35	1.8
	0.01	8.69	0.15	8.80	1.6
	0.02	8.70	0.30	8.89	1.9

[a] Total concentration of all species. [b] K_{N} is the equilibrium constant for formation of the complex from the tertiary amine base and the acidic species of the boron compound. The values of K_{N} were calculated from eq 8 in the text; the values of K_{BH} in water at 25° and 1 *M* ionic strength were taken from Table I; the values of $[\mathrm{RB(OH)_2}]_0$ and $[\mathrm{R\bar{B}(OH)_3}]_0$ were calculated from the pH of the buffer in the absence of amine and the total concentration of buffer by use of the equation in footnote a of Table 1; the value of $\mathrm{p}K_{\mathrm{R'XH}}$ for 1-methylimidazolium was determined to be 7.34 at 25° and 1 *M* ionic strength by potentiometric titration; and the value of $\mathrm{p}K_{\mathrm{R'XH}}$ for pyridinium at 25° and 1 *M* ionic strength with KCl is 5.52, according to W. P. Jencks and M. Gilchrist, *J. Amer. Chem. Soc.*, **90**, 2622 (1968).

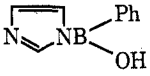
concentrations of 1-methylimidazole in place of [R'X]. Sample calculations in which we have assumed various possible values for [R'X] show that this approximation does not change the values of K_N for any of the cases given in Table II, including that of benzeneboronic acid plus pyridine, by more than 15%. The values of K_N for the reaction of 1-methylimidazole with benzeneboronic acid which have been obtained by the potentiometric method in this way are the same over a range of buffer ratios and 1-methylimidazole concentrations and agree well with the value from the spectrophotometric method (Table II).

The potentiometric method was also used to search for complex formation between benzeneboronic acid and pyridine (Table II). Small increases in pH occurred when pyridine was added to buffers of benzeneboronic acid. Since these increases may be partly due to other factors, such as a solvent effect upon pK_{HB} (see Table I), the small value for K_N is probably best considered only an upper limit. No complex formation with pyridine was detected spectrophotometrically (see below).

On the basis of the above evidence the most likely structures for the complexes between benzeneboronic acid and 1-methylimidazole or imidazole is **1**. *A priori,* an alternate possible structure for the complex with imidazole was **2**, the product of the dehydration of **1**. However, the fact that the equilibrium constant for the formation of the complex with imidazole is virtually the same as that for the formation of the complex with 1-methylimidazole and the fact that the ultraviolet spectra of the two complexes are the same shows that **2** is not formed in significant amounts. Philipp and Bender have reported B^{11} nmr evidence for the formation of a complex between imidazole and dimethyl *m*-nitrobenzeneboronate in methanol.[12]



1, R = CH_3 or H



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Search for Other Reactions.—On the basis of the known chemistry of boron compounds,[13-15] we expect that the following equilibria between oxygen, sulfur, and carbon nucleophiles and boric or benzeneboronic acids are rapidly established in water.

$$\begin{tikzcd}
R'X^- + RB(OH)_2 \arrow[r, shift left, "K_N"] \arrow[d, shift left] & R'X\bar{B}(OH)_2R \arrow[l, shift left] \arrow[d, shift left] \\
R'XH + RB(OH)_2 \arrow[r, shift left, "K_E"] \arrow[u, shift left, "K_{R'XH}\ \pm H^+"] & R'X\bar{B}(OH)R + H_2O \arrow[l, shift left] \arrow[u, shift left, "\pm H^+"']
\end{tikzcd} \qquad (9)$$

$$X = O, S, C$$

The expected equilibria for primary and secondary amines are somewhat more complicated.

$$\begin{tikzcd}
 & R_2'N\bar{B}(OH)_2R \arrow[d, shift left] \\
R_2'NH + RB(OH)_2 \arrow[r, shift left, "K_E"] & R_2'NB(OH)R + H_2O \arrow[l, shift left] \arrow[u, shift left, "\pm H^+"'] \arrow[d, shift left] \\
R_2'NH + RB(OH)_2 \arrow[r, shift left, "K_N"] \arrow[d, shift left] & R_2'H\overset{+}{N}\bar{B}(OH)_2R \arrow[l, shift left] \arrow[u, shift left] \arrow[d, shift left] \\
R_2'\overset{+}{N}H_2 + RB(OH)_2 \arrow[r, shift left, "K_X"] \arrow[u, shift left, "K_{R'XH}\ \pm H^+"] & R_2'H\overset{+}{N}B(OH)R + H_2O \arrow[l, shift left] \arrow[u, shift left, "\pm H^+"']
\end{tikzcd} \qquad (10)$$

In the case of tertiary amines, only the equilibria in eq 10 with constants K_N and K_X are possible.

A spectrophotometric search for these reactions was undertaken. The spectra of a number of nucleophiles which absorb strongly in the visible or ultraviolet region were determined at one pH value in the presence and absence of a great excess of boric acid, which itself shows only low end absorption (ϵ_{230} is less than 0.20 M^{-1} cm^{-1} for potassium borate buffer, 60% base) (Table III). Also, the spectrum of benzeneboronic acid alone was compared with its spectrum at the same pH value in the presence of an excess of each of several nucleophiles, after subtraction of the spectrum of the nucleophile alone (Table III). With the exception of the case of hydrazine and benzeneboronic acid at pH 8.35, the spectrum of the absorbing compound in each case was not altered significantly by the reagent in excess during the period of time given in Table III. In the presence of hydrazine buffer at pH 8.35 the spectrum of benzeneboronic acid shows λ_{max} 265 nm, ϵ 370 M^{-1} cm^{-1}, which is similar to that of benzeneboronic acid alone, but the spectrum lacks the weaker maxima at 259 and 272 nm which benzeneboronic acid alone shows and the molar absorbancy in the 265–300 nm range is greater. This difference may be due to the formation of the $H_2NNHB(OH)Ph$ species;[16] however, we have not investigated this reaction in detail.

Table III includes estimates of the upper limits for the values of K_N (see eq 9 and 10). These estimates were made by assuming that the spectrum of the tetrahedral adducts would be the same as that of the acidic species of the nucleophile in the case of the search for reactions with boric acid and the same as that of the benzeneboronate anion in the case of the search for reactions with benzeneboronic acid.

Upper limits for the equilibrium constants for the formation of the trigonal species, $R'XB(OH)R$ and $R_2'HN^+B(OH)R$ (eq 9 and 10) ,cannot be estimated from the spectra in acidic solution because we cannot estimate the spectral changes which would accompany the formation of these boron compounds. Although it seems likely that the spectra of these boron compounds would differ somewhat from those of the reactants, the absence of spectral changes is not definitive evidence for no reaction. It is worth noting that the absence of spectral changes with the absorbing nucleophiles and boric acid buffers at alkaline pH shows that significant amounts of the trigonal species $R'XB(OH)R$ and $R_2'HN^+B(OH)R$ have not formed under these conditions,

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[12] M. Philipp and M. L. Bender, *Proc. Nat. Acad. Sci. U. S.*, **68**, 478 (1971).

[13] E. L. Muetterties, Ed., "The Chemistry of Boron and Its Compounds," Wiley, New York, N. Y., 1967.

[14] H. Steinberg, "Organoboron Chemistry," Vol. I, Wiley, New York, N. Y., 1964.

[15] H. Steinberg and R. J. Botherton, "Organoboron Chemistry," Vol. II, Wiley, New York, N. Y., 1966.

[16] M. S. Bains, *Can. J. Chem.*, **44**, 534 (1966).

TABLE III

SPECTROPHOTOMETRIC SEARCH FOR REACTIONS BETWEEN NUCLEOPHILES AND BORIC OR BENZENEBORONIC ACID, IN WATER AT 25° AND 1.0 M IONIC STRENGTH

Registry no.	Nucleophile	pK^a	Concn, M	Boron compd, M	pH^b	Time^c interval, hr	$K_N{}^d$ M^{-1}
			Boric Acid				
62-53-3	Aniline	4.58	10^{-3}	1.0	9.4	12	0.15
			4.6×10^{-3}	0.32	1.0	12	
108-95-2	Phenol	9.95	3.6×10^{-4}	1.0	9.50	12	0.8^e
			9.1×10^{-4}	0.28	1.0	12	
100-02-7	*p*-Nitrophenol	7.14	7×10^{-5}	0.30	9.13	24	0.5
			7×10^{-5}	0.20	0.7	45	
108-98-5	Benzenethiol	6.5	1.5×10^{-5}	0.30	9.1	1	5.5
			1.5×10^{-5}	0.20	0.7	0.25	
1849-36-1	*p*-Nitrobenzenethiol	5.1	2.0×10^{-4}	0.30	9.1	1	0.7
10494-75-4	2-Methoxyethanethiol	9.5	4×10^{-4}	1.0	9.5	0.05^f	0.4
75-52-5	Nitromethane	10.3	1.2×10^{-3}	0.50	10.0	0.5^g	10
			Benzeneboronic Acid				
109-85-3	2-Methoxyethylamine	9.2	1.0	5.0×10^{-3}	9.08		5
			1.0	5.0×10^{-3}	8.68	22	4^h
			0.90	3.0×10^{-3}	1.0		
280-57-9	Triethylenediamine	2.9, 8.8	0.80	3.0×10^{-3}	8.49		2.2
			0.80	1.8×10^{-3}	5.2		0.3
302-01-2	Hydrazine	8.1	0.50	3.0×10^{-3}	8.35		*i*
			0.50	3.0×10^{-3}	1.3		0.25
110-91-8	Morpholine	8.3	0.70	3.0×10^{-3}	8.58		2.0
			0.70	3.0×10^{-3}	1.0		
583-58-4	3,4-Lutidine	6.5	0.052	0.0158	7.62		1.6^j
110-86-1	Pyridine	5.2	0.08	1.1×10^{-3}	5.2		5
753-90-2	2,2,2-Trifluoroethylamine	5.7	0.4	3.3×10^{-3}	5.7		0.5
57-12-5	Cyanide	9.4	1.0	4.8×10^{-3}	8.60	*k*	1.3
16887-00-6	Chloride		1.0	3.0×10^{-3}	2.7		0.12^l

[a] Of conjugate acid; from W. P. Jencks in "Handbook of Biochemistry," H. A. Sober, Ed., Chemical Rubber Publishing Co., Cleveland, Ohio, 1968, pp J150–J189. [b] Adjusted with KOH or HCl. [c] Period between preparation of the reaction mixture and the last recording of its spectrum. Where no time is noted, the spectrum was recorded once about 15 min after preparation of the reaction mixture. [d] These values for K_N are upper limits. In the case of the boric acid buffers, they have been calculated by use of the equation

$$K_N = \frac{\Delta A(a_H + K_{R'XH})}{([R'XH]_T\Delta\epsilon - \Delta A)K_{R'XH}/[RB(OH)_2]}$$

where ΔA is the difference between the absorbance of the nucleophile in the absence of boric acid buffer and that in the presence of boric acid buffer, $[R'XH]_T$ is the total concentration of the nucleophile in all species, $\Delta\epsilon$ is the difference between the apparent molar absorbancy of the nucleophile at the pH of the reaction mixture and the molar absorbancy of the acidic species alone, $K_{R'XH}$ is the acid dissociation constant of the nucleophile, and $[RB(OH)_2]$ is the concentration of boric acid, including polymeric species, in the boric acid buffer. In the case of benzeneboronic acid, the upper limits for the values of K_N were calculated in a similar manner by use of eq 3 and 4 in the text, except that the concentration of the basic species of the nucleophile was known from the preparation of the buffer. The calculations were made by taking ΔA as 0.1, usually at the wavelength where $\Delta\epsilon$ is maximal; in fact, ΔA was less than 0.05 except in the cases which are noted. [e] The spectra showed ΔA at 233 nm, where $\Delta\epsilon$ is 2,800 M^{-1} cm^{-1}, to be about 0.10 (1-cm path length). This small ΔA may be due either to a conversion of about 10% of the phenol to a boron adduct or to an increase in the apparent pK of phenol by 0.06 log units. [f] The absorbance of the reaction mixtures with and without borate buffer decreased after this time, probably because of oxidation of the thiol to the disulfide. [g] The spectrum did not change between 20 and 30 min after preparation of the solution. Calculation based upon the rate constants for ionization of nitromethane [R. P. Bell and D. M. Goodall, *Proc. Roy. Soc., Ser. A*, **294**, 273 (1966)] shows that about 10 min is sufficient time for equilibration of the acid and anion. [h] ΔA at 267 nm, where the $\Delta\epsilon$ term of eq 3 is 155 M^{-1} cm^{-1}, was 0.12 (1-cm light path). [i] A reaction occurs. See the text. [j] Ionic strength, 0.53 M. [k] Cyanide buffer slowly forms a compound which absorbs at 293 nm. For this reason the reaction mixtures which contained cyanide buffer were prepared by using stock solutions of KCN and HCl, and the spectra of such mixtures with and without benzeneboronic acid present were determined at the same time after preparation. [l] The reaction mixture without chloride had an ionic strength of 0.002 M.

since such reaction would decrease the total concentration of the free nucleophile and thus decrease the concentration of the basic species and the absorbance at wavelengths where the basic species absorbs strongly. The trigonal species, $R_2'NB(OH)R$, which can form from primary and secondary amines, would also not have been detected if its spectrum were identical with that of $R_2'NH$ or $RB(OH)_2$. Moreover, if a large fraction of the absorbing compound were undetected $R_2'NB(OH)R$, then the values of the upper limits for K_N given in Table III would be incorrect.

Comparisons of Reactivities.—The affinities of imidazole and 1-methylimidazole for benzeneboronic acid are much higher than we would expect on the basis of the data for other nitrogenous bases given in Table III and their affinities for the proton (the pK's of imidazole and 1-methylimidazole are 7.22 and 7.34). Two possible explanations for this unusual affinity are that a positive charge next to tetrahedral boron is unfavorable and delocalization of this charge in the imidazole adducts (see **1**) is thus relatively stabilizing and/or that there is a strong electrostatic interaction between the positively charged imidazolium nucleus and the partially negative phenyl ring in the adducts. In order to decide between these two possibilities, it will be necessary to examine the affinities of a series of nitrogenous bases for an alkylboronic acid.

Finally, it is interesting to note that our results show that the affinities of nitrogen, sulfur, and carbon nucleophiles for trigonal boron are not much greater than that of an oxygen nucleophile of the same basicity. This finding is different from that for a carbonyl carbon, for which nitrogen, sulfur, and carbon nucleophiles have a greater affinity than an oxygen nucleophile of the same basicity.[17] Thus, in this sense trigonal boron is more similar to the proton than to trigonal carbon in its chemistry.

Registry No.—Boric acid, 10043-35-3; benzeneboronic acid, 98-80-6; 1-methylimidazole, 616-47-7.

[17] E. G. Sander and W. P. Jencks, *J. Amer. Chem. Soc.*, **90**, 6154 (1968).

A Direct ^{1}H and ^{19}F Nuclear Magnetic Resonance Study of Boron Trifluoride Complexes with Acetylacetone, Methoxyacetonitrile, Methoxypropionitrile, Methylacetoacetate, Methylmethoxyacetate, Methoxymethylacetate, 2-Methoxyethylacetate, and Methylpyruvate

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A direct proton and fluorine-19 nuclear magnetic resonance chemical shift and area study of boron trifluoride complexes with acetylacetone, methylacetoacetate, methoxyacetonitrile, methoxypropionitrile, methylmethoxyacetate, methoxymethylacetate, 2-methoxyethylacetate, and methylpyruvate has been completed. In these systems ligand exchange is slow enough below −50° to permit the direct observation of pmr signals for bulk ligand and molecules bound to the boron trifluoride. In the nitrile solutions the ^{1}H and ^{19}F nmr spectra indicate complete complexing at the methoxy linkage. The most stable complexes in solutions of acetylacetone and methylacetoacetate involve the enol tautomers of these molecules. In the remaining systems, complexing at more than one site in the base is indicated by the ^{1}H and ^{19}F nmr data. The competition by each oxygen site is influenced by basicity differences, steric hindrance, and resonance.

Studies of boron trihalide complexes with organic bases have been undertaken by several calorimetric[1-4] and spectroscopic[5-11] techniques to ascertain the chemical and structural features of the components which influence these interactions. Recent measurements have demonstrated the utility of the direct nuclear magnetic resonance method as a supplementary tool for these investigations.[12-16] The success of this method is based on the ability to slow ligand exchange, thereby allowing the observation of separate nmr signals for bulk ligand and ligand bound to the boron trihalide. When this observation is possible, quantitative measurements of the chemical shifts induced by complex formation, the stoichiometry of the complex, the ligand interaction site, steric effects, and ligand preference can be made. A correlation of these results with those obtained by other techniques frequently is possible.

The present study involves a series of ligands which contain more than one possible interaction site, an oxygen atom in most cases. It was hoped that a measure of competitive complex formation would be possible with several of these molecules. If so, the influence of properties such as molecular complexity, the relative basic strengths of the functional groups present, tautomerism, and molecular resonance structures could be ascertained.

Experimental Section

Methods.—All organic chemicals were of the highest commercial grade available, and they were distilled before use. The boron trifluoride was fractionated at −110°. The samples were prepared *in vacuo* and the nmr tube was sealed and stored in liquid nitrogen until the spectrum could be recorded. Each sample contained a few per cent by volume of tetramethylsilane (TMS) and hexafluorobenzene for use as internal nmr chemical shift standards for ^{1}H and ^{19}F nuclei, respectively.

The chemical shift and area measurements were made with a Varian HA-100 spectrometer, operating at 94.1 MHz for the ^{19}F experiments. With each sample, the nmr measurements for each nucleus were repeated at least a day later to ensure that decomposition had not occurred. This was not a problem with any system reported here. The procedure has been described in more detail in previous publications and it involves observing the nmr spectrum as the sample is cooled in the spectrometer probe.[12-16] In most of the systems studied, exchange was slow enough only at temperatures below −50° to permit the observation of separate bulk and bound ligand signals.

Results.—A summary of all pmr chemical shift and integration data is presented in Table I. As mentioned in the table, the concentrations in all cases were maintained at mole ratios of 1:3:60, BF_3 to base to dichloromethane. Such dilutions were used to avoid intermolecular interactions between ligand molecules. Thus, the chemical shift separations between bound and bulk ligand molecules, represented by the quantity $\Delta\nu(C - B)$ in Table I, are an accurate measure of the effect of complex formation by BF_3 with these bases.

[1] H. C. Brown and R. R. Holmes, *J. Amer. Chem. Soc.*, **78**, 2173 (1956).

[2] T. D. Coyle and F. G. A. Stone, *ibid.*, **83**, 4138 (1961).

[3] M. F. Lappert and J. K. Smith, *J. Chem. Soc.*, 7102 (1965).

[4] N. N. Greenwood and T. S. Srivastava, *J. Chem. Soc. A*, 267, 270 (1965).

[5] J. M. Miller and M. Onyszchuk, *Can. J. Chem.*, **42**, 1518 (1954).

[6] E. Gore and S. S. Danyluk, *J. Phys. Chem.*, **69**, 89 (1964).

[7] M. Okada, K. Suyama, and Y. Yamashita, *Tetrahedron Lett.*, 2329 (1965).

[8] P. N. Gates, E. J. McLauchlan, and E. F. Mooney, *Spectrochim. Acta*, **21**, 1445 (1965).

[9] S. J. Kuhn and J. S. McIntyre, *Can. J. Chem.*, **43**, 375 (1964).

[10] N. N. Greenwood and B. H. Robinson, *J. Chem. Soc. A*, 511 (1966).

[11] R. J. Gillespie and J. S. Hartman, *Can. J. Chem.*, **45**, 859 (1966).

[12] A. Fratiello, T. P. Onak, and R. E. Schuster, *J. Amer. Chem. Soc.*, **90**, 1194 (1968).

[13] A. Fratiello and R. E. Schuster, *Inorg. Chem.*, **7**, 1581 (1968).

[14] A. Fratiello and R. E. Schuster, *Org. Magn. Resonance*, **1**, 139 (1969).

[15] A. Fratiello and R. E. Schuster, *Inorg. Chem.*, **8**, 480 (1969).

[16] A. Fratiello, R. E. Schuster, and M. Geisel, *ibid.*, **11**, 11 (1972).