Table I. Typical Results of the Quantitative Determination of Adipic Acid by Gas Chromatographic Analysis

Actual wt %	Ratio peak areas ± std dev[a]	Wt % by GC	Relative error, %	Solvent[b]
7.00[c]	0.172 ± 0.001	6.95	0.7	DMF
7.50	0.207 ± 0.004	7.75	3.3	DMF
10.00	0.299 ± 0.001	9.85	1.5	DMF
12.00[c]	0.397 ± 0.008	12.05	0.4	DMF
17.0[c]	0.603 ± 0.010	16.80	1.2	DMF
18.0	0.643 ± 0.010	17.75	1.4	DMF
20.0[c]	0.756 ± 0.015	20.2	1.0	DMF
21.0	0.797 ± 0.021	21.25	1.2	DMF
25.0[c]	0.961 ± 0.020	25.05	0.2	DMF
5.00	0.081	4.90	2.0	EtOH
10.00	0.240	9.70	3.0	EtOH
15.0	0.433	15.50	3.3	EtOH
20.0	0.600	20.5	2.5	MeOH
5.90	0.120	6.10	3.3	HOAc
13.30	0.380	13.9	4.5	HOAc

[a] The average of from 3 to 5 runs.
[b] Different calibration curves were required for each solvent and, in each case, straight lines were obtained from 4 or 5 different concentrations.
[c] Samples were used to calibrate the analytical method in DMF.

essary to exclude the possibility of dehydration of adipic acid to adipic anhydride on the column because the parent peak for adipic acid, *m/e* 146, was not observed in the mass spectrum.

A calibration curve representing weight per cent of adipic acid *vs.* the ratio (area of adipic acid peak to area of internal standard peak) was determined from the data reported in Table I. The curve is linear throughout the range 7 to 25 wt % of adipic acid. The curve is nonlinear below 7 wt % adipic acid and the extrapolated intercept is only apparent. The technique is applicable to analyses of samples containing less than 7% adipic acid by direct comparison to the nonlinear portion of the calibration curve. The accuracy and precision of the analytical technique are demonstrated by the results summarized in Table I.

Succinic and glutaric acids were also eluted from this column. However, under the chromatographic conditions used for adipic acid analysis, these compounds tailed considerably and no attempt was made to optimize the elution and resolution of these acids.

Adipic acid can also be eluted employing an untreated Porapak Q column or a 5% FFAP–Chromosorb W column; however, extensive tailing is observed. This difficulty was minimized when the 7% FFAP–Porapak Q column was used.

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Consecutive Titrimetric Determination of Boron and Nitrogen in Amineboranes

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THE DETERMINATION of boron in numerous compounds through the formation of boric acid and its subsequent titration with aqueous base in the presence of mannitol or glycerol is well known (*1*). Difficulties have been reported, however, in determining this element in the presence of nitrogen in the analysis of various boron–nitrogen compounds. In 1963, Yoshizaki (*2*) described the development of a flame photometric method for determining boron in organoboron compounds, which he applied to the analysis of several borazines, in view of "ambiguities such as starting and end points of titration" which he described as characteristic of the determination of boron in compounds containing nitrogen. Procedures for the separation of boron from nitrogen in the analysis of various amineboranes also have been described. One of the more common of these involves separation, by distillation, of a volatile compound such as trimethylborate which is formed by the reaction of the amineborane with acidified methanol (*3*, *4*). The borate is subsequently hydrolyzed and the hydrolyzate titrated as described above. Nitrogen is determined, frequently, by the Kjeldahl or Dumas method on a separate sample of the amineborane.

In reference to this problem, it seems desirable to report that, for a wide variety of amineboranes, boron and nitrogen can be determined in the presence of each other by a straightforward titrimetric procedure which eliminates the time consuming separation of these elements and a separate nitrogen determination. It is necessary only to hydrolyze the amineborane and obtain a soluble hydrolyzate in which the amine fragment has not been degraded. Consecutive titrations with standard base are carried out in aqueous solution with the use of a pH meter, boron being determined by a conventional titration in the presence of mannitol, and nitrogen by titration of the corresponding ammonium ion. The method is illustrated here for the analysis of a group of amineboranes showing wide variations in structure.

EXPERIMENTAL

Hydrolysis. All samples were hydrolyzed in aqueous hydrochloric acid. Because of their greater kinetic stability, it was necessary to heat gently solutions of quinuclidineborane and triethylenediamine bisborane with a bunsen flame to obtain complete hydrolysis. Hydridic hydrogen was determined by measurement of the evolved hydrogen.

Titration of the Hydrolyzate. Each hydrolyzate was adjusted to pH = 2–3 with concentrated sodium hydroxide,

(1) N. H. Furman, Ed., "Scott's Standard Methods of Chemical Analysis," 6th ed., Vol. I, Van Nostrand, Princeton, 1962, Chap. 8.
(2) T. Yoshizaki, ANAL. CHEM., **35**, 2177 (1963).
(3) H. Nöth and H. Beyer, *Ber.*, **93**, 928 (1960).
(4) G. E. Ryschkewitsch and E. R. Birnbaum, *Inorg. Chem.*, **4**, 575 (1965).

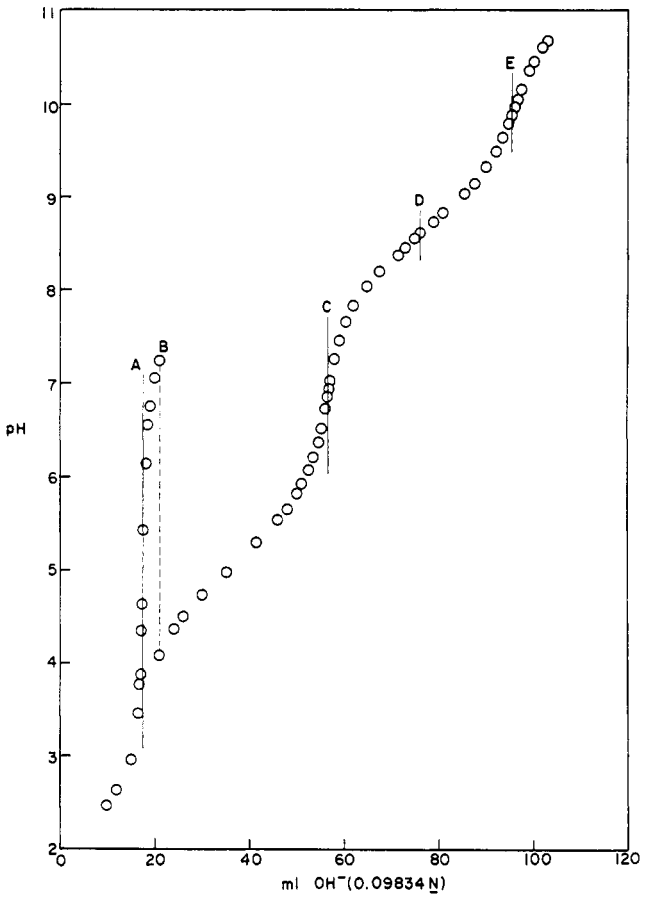


Figure 1. Titration of morpholineborane hydrolyzate

Amineborane sample wt = 0.3874 g (3.837 mmole); A = 17.33 ml; 13 g mannitol added at B; C = 56.32 ml; $D = (E-C)/2$ = 75.96 ml, pH = 8.60; E = 95.60 ml; boron determination, $(C-A)$ = 3.834 mmole OH^-; nitrogen determination, $(E-C)$ = 3.863 mmole OH^-

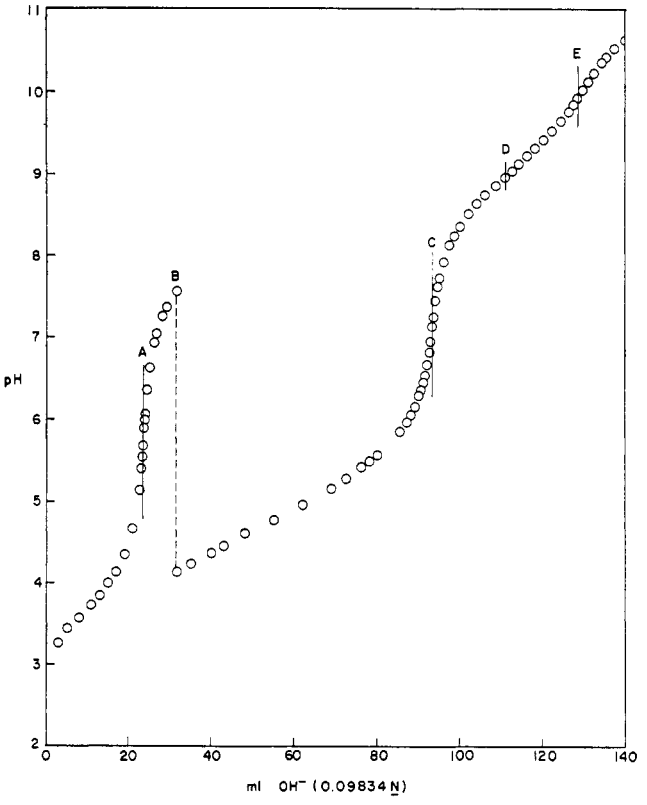


Figure 2. Titration of triethylenediamine bisborane hydrolyzate

Amineborane sample wt = 0.4809 g (3.438 mmole); A = 23.80 ml; 20 g mannitol added at B; C = 93.60 ml; $D = (E-C)/2$ = 111.0 ml, pH = 8.98; E = 128.4 ml; boron determination, $(C-A)$ = 6.864 mmole OH^-; nitrogen determination, $2(E-C)$ = 2(3.422) = 6.844 mmole OH^-

then titrated with standard sodium hydroxide through the first equivalence point at room temperature (25°–30° C) using a Beckman Model GS or Expandomatic pH Meter. Subsequent treatment of the solution was based on the acid strength of the amine salt. For solutions in which the pK_a of the ammonium ion exceeded about 7, the boron analysis preceded that of the amine salt. Mannitol (about 3–4 grams per 0.1-gram sample of amineborane) was added to complex the borate, and boron determined by titration to the next end point. Further titration allowed determination of the amine salt and estimation of its pK_a from the flat portion of the curve. When the pK_a of the ammonium ion was less than about 7, the sequence was reversed and titration of the amine salt preceded the boron analysis.

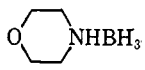
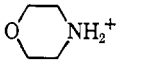
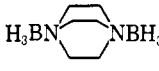
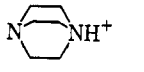
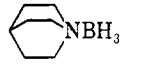
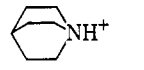
RESULTS AND CONCLUSIONS

Data obtained in the analysis of several amineboranes are given in Table I. Figures 1 and 2 are titration curves for the analysis of compounds in which the resulting amine salts are relatively weakly acidic, while Figures 3 and 4 represent cases in which more strongly acidic ammonium ions are produced. The amounts of base required for determination of boron and nitrogen in individual amineboranes are given in the captions in terms of millimoles of hydroxide and amineborane.

For the titration of ammonium ions having pK_a values greater than about 9.5, points of inflection are slight and, consequently, the accuracy of the nitrogen determination is diminished. Amine salts having pK_a values less than about 4, are sufficiently acidic to be titrated with the excess hydronium ion introduced as hydrochloric acid in the hydrolysis of the amineborane. Nevertheless, the method may be used in the analysis of diamine bisboranes, where hydrolysis produces a diammonium ion, even when only one of the corresponding acid functions can be titrated conveniently. For example, in the previously reported analysis of ethylenediamine bisborane (ethane 1,2-diamineborane) (*5*), the end point in the neutralization of the first $—NH_3^+$ group of the ethylenediammonium ion is obtained readily, but the β-aminoethylammonium ion is too weakly acidic (pK_a = 9.93) (*6*) to be determined accurately. For *m*- and *p*-phenylenediamine bisboranes and for triethylenediamine bisborane, the first acid dissociation constants of the respective diammonium ions are high and a separation of the first ammonium ion from excess hydronium ion is not obtained. The second dissociation constants, however, are in a range amenable for titration of the corresponding monoammonium ions. The aminophenylammonium ions are relatively acidic and, therefore, are titrated prior to the determination of boron (Figure 3) while the reverse is true for the monoammonium ion of triethylenediamine (Figure 2). The determination of only one of the acid functions in a diammonium ion (and estimation of the pK_a for that ion) is usually sufficient for characterization and analysis of the bisborane.

(5) H. C. Kelly and J. O. Edwards, *J. Am. Chem. Soc.*, **82**, 4842 (1960).
(6) "Handbook of Chemistry," N. A. Lange, Ed., Handbook Publishers, Inc., Sandusky, Ohio, 1946.

Table I. Analysis of Amineboranes

Amineborane	R_3NH^+	Calcd. pK_a[a]	Calcd. H^-	Calcd. B	Calcd. N	Found pK_a	Found H^-	Found B	Found N
$H_3BH_2NCH_2CH_2NH_2BH_3$ (*5*)	$H_3NCH_2CH_2NH_3^{2+}$	7.31[b]	6.89	24.6	31.9	7.31	6.82	24.3	32.0
		8.7 (*8*)	3.00	10.72	13.87	8.60	2.97	10.71	13.97
		8.98[c]	4.32	15.47	20.03	8.98	4.30	15.44	19.94
$HOCH_2CH_2NH_2BH_3$ (*9*)	$HOCH_2CH_2NH_3^+$	9.44[d]	4.04	14.44	18.69	9.6	4.00	14.15	18.32
		10.58 (*11*)	2.42	8.65	11.21		2.34	8.63	
p-$CH_3C_6H_4NH_2BH_3$	p-$CH_3C_6H_4NH_3^+$	5.30[e]	2.50	8.94	11.58	5.22	2.40	8.98[f]	11.78
m-$CH_3C_6H_4NH_2BH_3$	m-$CH_3C_6H_4NH_3^+$	4.74[e]	2.50	8.94	11.58	4.8	2.39	9.07[f]	11.36
m-$C_6H_4(NH_2BH_3)_2$	m-$H_2NC_6H_4NH_3^+$	5.0[c]	4.45	15.9	20.6	5.04	4.34	15.5	20.5
p-$C_6H_4(NH_2BH_3)_2$	p-$H_2NC_6H_4NH_3^+$	6.23[c] 6.04[e]	4.45	15.9	20.6	6.18	4.20	15.3	20.1
$CH_3NH_2BH_3$	$CH_3NH_3^+$	10.64[d]	6.73	24.10	31.20		6.61	23.58	
$(CH_3)_3CNH_2BH_3$	$(CH_3)_3CNH_3^+$	10.45[d]	3.48	12.44	16.10		3.45	12.7	

[a] Lit. value at 25° C.
[b] At 30° C, 0.2*M* HCl, 0.1*M* $BaCl_2$, 0.1*M* KCl (*7*).
[c] Obtained from titration of separate sample of the diamine dihydrochloride.
[d] Calculated from pK_b (*10*).
[e] Calculated from pK_b (*6*).
[f] No mannitol added.

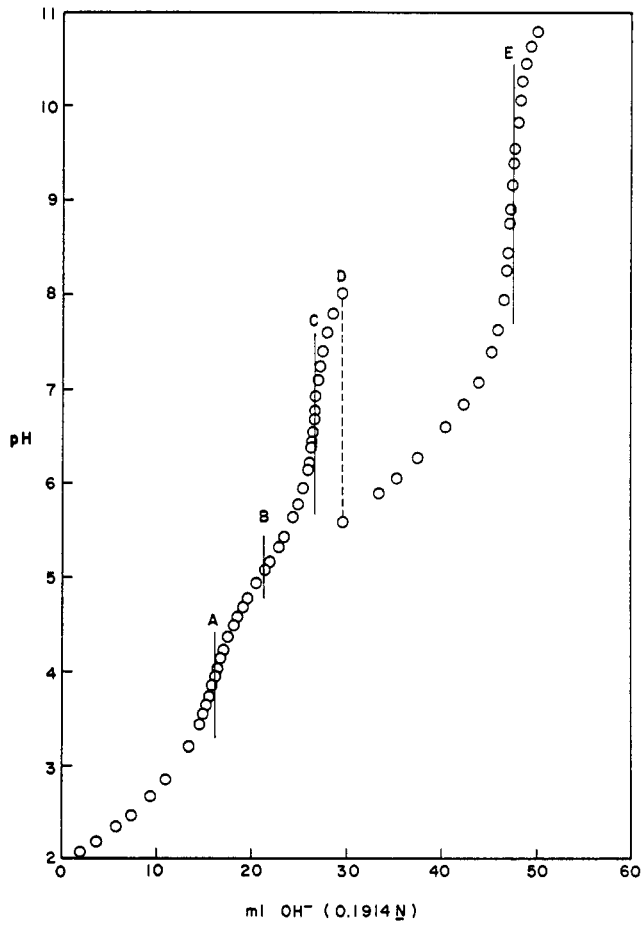


Figure 3. Titration of *m*-phenylenediamine bisborane hydrolyzate

Amineborane sample wt = 0.2776 g (2.044 mmole); *A* = 16.11 ml; *B* = (*C*–*A*)/2 = 21.41 ml. pH = 5.04; *C* = 26.71 ml; about 15 g mannitol added at *D*; *E* = 47.45 ml; nitrogen determination, 2-(*C*–*A*) = 2(2.029) = 4.058 mmoles OH^-; boron determination, (*E*–*C*) = 3.970 mmoles OH^-

(7) J. Bjerrum, "Metal Ammine Formation in Aqueous Solution, Theory of Reversible Step Reactions," P. Haase and Son, Copenhagen, 1941, p. 206.
(8) R. J. Bruehlman and F. H. Verhoek, *J. Am. Chem. Soc.*, **70,** 1401 (1948).
(9) H. C. Kelly and J. O. Edwards, *Inorg. Chem.*, **2,** 39 (1963).
(10) "Handbook of Chemistry and Physics," 47th ed, The Chemical Rubber Co., Cleveland, Ohio, 1966.
(11) B. M. Wepster, *Rec. Trav. Chim.*, **71,** 1171 (1952).

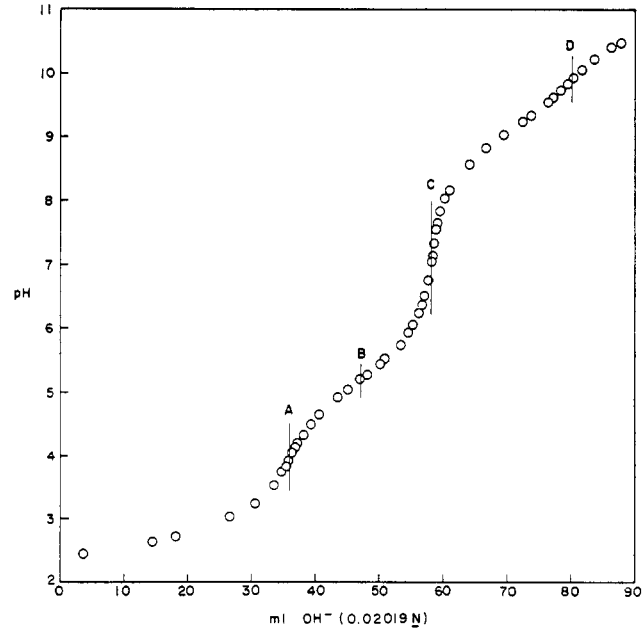


Figure 4. Titration of *p*-toluidineborane hydrolyzate

Amineborane sample wt = 0.0533 g (0.440 mmole); *A* = 36.00 ml; *B* = (*C*–*A*)/2 = 47.10 ml, pH = 5.22; *C* = 58.20 ml; *D* = 80.10 ml; no mannitol added during analysis. Midpoint between *C* and *D* = 69.15 ml, pH = 9.05. Nitrogen determination, (*C*–*A*) = 0.448 mmole OH^-; boron determination, (*D*–*C*) = 0.442 mmole OH^-

Even when the acidity of the amine salt originating from hydrolysis of a monoamineborane is outside the range for accurate determination, the direct analysis of boron in the presence of the amine salt (or amine) presents little or no difficulty. The determination of boron in quinuclidineborane, as well as methyl and *t*-butylamineboranes are illustrative (Table I).

It is evident from Table I that somewhat better agreement between experimental and theoretical values are realized with the tertiary alkyl- and heterocyclic amineboranes than with the primary alkyl- and arylamineboranes. For the most part, this is a reflection of the state of purity in which these compounds were prepared, the latter compounds being less stable kinetically and more difficult to prepare and maintain in high purity. This analytical method, nevertheless, is useful for amineboranes which exhibit a wide range of hydrolytic stability. It has been applied to compounds which could be hydrolyzed in aqueous hydrochloric acid. Applications to systems where stronger oxidants are required for conversion of the boron-containing species to borate, and to nonaqueous or mixed solvent systems has not, as yet, been explored.

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A Technique for Increasing the Solvent Properties in Proton Magnetic Resonance Spectrometry

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THE USE of $AsCl_3$, $SbCl_3$, and mixtures of these two halides as solvents for PMR and infrared spectrometry has been previously reported (*1*, *2*). Others have also found these solvents useful for dissolving various compounds in order to obtain PMR spectra (*3*).

The present note involves using these halides in combination with other solvents. We have encountered some compounds that are slightly soluble in PMR solvents such as deuterated chloroform, carbon disulfide, and carbon tetrachloride. We found that the addition of approximately 5% $SbCl_3$ to these solvents increased the solubility of the compounds with only a small shift in their PMR signals.

We have chosen for examination three classes of compounds. While all the compounds in these classes are drugs we have also examined nondrugs. These latter did not fall into a series of structurally related compounds and so are not reported here. We assume that our solubility data could be extended to any compounds related to the ones we report whether they are drugs or not.

The three classes of compounds examined were substituted barbituric acids, steroids, and sulfa derivatives. It was found that the substituted barbituric acids, mephobarbital, secobarbital, pentobarbital, phenobarbital, and the related compound diphenylhydantoin (5,5-diphenyl-2,4-imidazolidinedione) have enhanced solubility in $SbCl_3$–$CDCl_3$. In addition secobarbital, pentobarbital, and phenobarbital show enhanced solubility in CS_2–$SbCl_3$. The steroids, estrone (3-hydroxyestra-1,3,5(10)-trien-17-one) and prednisolone (11β,-17,21-trihydroxypregna-1,4-diene-3,20-dione) also have enhanced solubility in $CDCl_3$–$SbCl_3$. $SbCl_3$ is not soluble in CCl_4 so that we made no studies of this solvent system. The addition of $AsCl_3$ to CCl_4 increases the solubility of the compounds over that observed in CCl_4 alone. However, the toxic nature of $AsCl_3$ makes the $SbCl_3$–$CDCl_3$ a more favorable solvent.

Inasmuch as $SbCl_3$ is acidic in nature, mixed solvents containing this halide cannot be used to solubilize most basic compounds. Most basic compounds react with halide to form insoluble hydrochlorides.

The combination of $SbCl_3 + CHCl_3$ has been used as a detection agent for TLC. It has been found that many organic compounds which do not fluoresce before treatment fluoresce after being sprayed with this mixture. This indicates that probably a charge transfer complex is formed between the organic compound and $SbCl_3$. This type of interaction suggests that the increased solubility we obtain with the mixed solvents is due to a charge transfer complex formed by the drug and $SbCl_3$.

There is about a 0.15-ppm downfield shift of signals in the mixed solvent as compared to the signals in CS_2 or $CDCl_3$ which is probably caused by the diamagnetic properties of the mixed solvent as well as the formation of a complex (*4*).

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(1) H. A. Szymanski and A. Bluemle, *J. Polymer Sci.*, **3**, 63 (1965).
(2) H. A. Szymanski, A. Bluemle, and W. Collins, *Appl. Spectry.*, **19**, 137 (1965).
(3) E. G. Brame, Jr., R. G. Ferguson, and G. J. Thomas, Jr., ANAL. CHEM., **39**, 517 (1967).
(4) H. A. Szymanski and R. Yelin, *J. Polymer Sci.*, in press.