TABLE II

PROPERTIES, YIELDS AND ANALYSIS OF METHYL ESTERS

Acid	A (Fischer) Yield,[a] %	M.p., °C.	B (Normal) Yield,[a] %	M.p., °C.	C (Pseudo) Yield,[a] %	M.p., °C.	Formula	Calcd. C	Calcd. H	Found[c] C	Found[c] H
VI	94	42.4–45.2	64	43.5–46.0	81	56–58	$C_{18}H_{18}O_3$	76.6	6.4	76.7	6.4n
										76.4	6.5n
										76.4	6.6p
VII	96	58.5–61.0	86	60.0–61.5	36	95.5–98.5	$C_{18}H_{18}O_3$	76.6	6.4	76.2	6.3n
										76.5	6.5p
X	4	111.8–113.6	94	112–113.6	46[b]	111–113	$C_{19}H_{20}O_3$	77.0	6.8	77.0	6.5n
										76.9	6.4n
XI	87	113–115	86	113–114.8	25[b]	113–115.2	$C_{19}H_{20}O_3$	77.0	6.8	77.1	6.7n
										77.3	6.5n

[a] In all cases, the per cent. yield means the yield of crystalline product with the indicated melting point. [b] These esters were the same as those prepared with diazomethane. A large amount of lower melting material was also obtained some of which was probably pseudo ester. [c] Analyses marked n are for normal esters, analyses marked p are for pseudo esters. Microanalyses by Mrs. E. H. Klotz.

man IR-2 instrument) in the ranges of 30, 35 and 40% pseudo ester working curves were obtained. The normal ester absorbed strongly at 6.00 and 7.95 μ and the pseudo weakly whereas the reverse was true for the other two points. By determining the extinction at all four positions for the ester mixture obtained by Fischer–Speier esterification of II it was determined that the mixture contained 35.5 ± 2.5% pseudo methyl ester. In a similar way, the composition of the esters formed from IX was found to be 32 ± 2% pseudo methyl ester. In the latter case, a Baird Associates instrument was used and the extinctions at 5.70 μ (high extinction for pseudo, low for normal), 5.80 and 6.00 (high extinction for normal, low for pseudo) were used for analysis. All measurements involved 10 mg. of ester per ml. of carbon disulfide and the same cell was used throughout an analysis for one pair of isomers.

COLUMBUS 10, OHIO RECEIVED MARCH 28, 1951

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY OF THE UNIVERSITY OF NEW HAMPSHIRE]

Kinetics of the Brominolysis of Benzeneboronic Acid in Aqueous Acetic Acid Solution

BY HENRY G. KUIVILA AND ELIOT K. EASTERBROOK

The kinetics of the brominolysis of benzeneboronic acid in aqueous acetic acid solution have been investigated. The reaction is first order in bromine and in benzeneboronic acid. Its rate increases with the water content of the solvent. The reaction shows a positive salt effect, catalysis by bases and retardation by acids. The bearing of these observations on the mechanism is discussed.

Electrophilic substitution reactions in the aromatic series, such as nitration and halogenation, have been frequently investigated and are fairly well understood. But similar reactions in which an atom or group other than hydrogen is replaced have, for the most part, been neglected from the standpoint of detailed kinetic investigation. An obvious source of compounds of this type is the organometallic field. However, only a few types of organometallic compounds have been found to undergo electrophilic displacement reactions at measurable rates. Among these the cleavage of organomercurials with acids have received considerable attention from Kharasch and his coworkers from the standpoint of reactivity.[1] More recently Corwin and Naylor[2] have investigated the kinetics of the cleavage of diphenylmercury by acetic and formic acids and Winstein[3] has discussed investigations on the reactions of organomercuric halides with halogens.

The aromatic boronic acids constitute another group of organometallic compounds in which some electrophilic displacement reactions proceed at measurable rates. Melnikoff[4] has shown that the reaction between bromine and aromatic boronic acids in water proceeds quantitatively

$$ArB(OH)_2 + Br_2 + H_2O \longrightarrow ArBr + H_3BO_3 + HBr \quad (I)$$

This reaction was chosen as the first of a series to be subjected to kinetic study. It was hoped thereby to gain some information concerning the mechanism of the reaction in particular and of electrophilic displacement reactions in general.

Experimental

Reagents.—The acetic acid was purified by treatment with chromic acid and distillation through a column with 40 theoretical plates. The acetic acid percentages given are percentages by volume. Sodium chloroacetate and sodium dichloroacetate were prepared from the acids by the method of Bischoff and Walden.[5] Sodium perchlorate (Eimer and Amend) was recrystallized from water and dried for several days at 110°. Baker C.P. sodium acetate was used in the experiments here reported. The Reagent Grade salt from another source gave anomalous kinetics. The sodium benzoate was Mallinckrodt U.S.P. and the other reagents were C.P. or Reagent Grade and were used without further purification.

The bromine was distilled from potassium bromide, the

---

(1) M. S. Kharasch, R. R. Legault and W. R. Sprowls, *J. Org. Chem.*, **3**, 409 (1938–1939), and preceding papers of the series.

(2) A. H. Corwin and M. A. Naylor, THIS JOURNAL, **69**, 1004 (1947).

(3) S. Winstein, Paper presented at the Conference on Reaction Mechanisms, Northwestern University, Evanston, Illinois, September 1, 1950.

(4) N. N. Melnikoff, *J. Gen. Chem. U. S. S. R.*, **6**, 636 (1936); *C. A.*, **30**, 5571 (1936); *ibid.*, **8**, 1766 (1938); *C. A.*, **33**, 4908 (1939).

(5) C. A. Bischoff and P. Walden, *Ann.*, **279**, 46 (1894).

middle half being retained for use. Benzeneboronic acid, prepared by the method of Bean and Johnson,[6] was crystallized from water at least twice and stored in a brown tightly-capped bottle. The samples used melted at 214-215° (uncor.).

Kinetic Experiments.—Appropriate volumes of solutions of the reagents except bromine were transferred to the reaction flask and brought to temperature in a bath maintained at 25.0 ± 0.03°. The bromine solution, also at 25.0°, was added at zero time. Aliquots (usually 5 or 10 ml.) were removed by pipet, transferred to a flask containing 5 ml. of 10% potassium iodide solution and titrated with thiosulfate. It was found experimentally that zero time titers could be determined accurately by extrapolation and this was the method used in most of the faster runs. The extrapolation usually covered less than one per cent. of the initial bromine titer. Only about 60% of the reaction mixture was used for kinetic points. As a result volatility losses were relatively small. Each result reported represents the mean of at least two runs which agreed to within 5%. Sodium perchlorate was used to adjust ionic strengths.

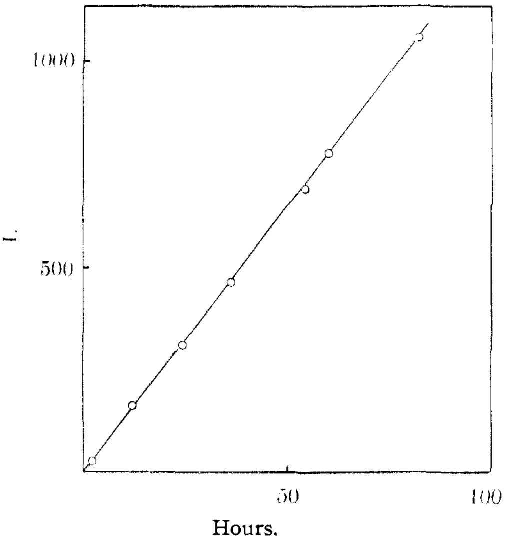


Fig. 1.—Kinetic plot for brominolysis of benzeneboronic acid in "50%" acetic acid at 25.0°; $[Br_2]_0$, 0.00570 M, $[C_6H_5B(OH)_2]_0$, 0.00994, $[Br^-]_0$, 0.25 M. Last point corresponds to 70% consumption of bromine; k = 0.00716 l. mole^{-1} sec.$^{-1}$.

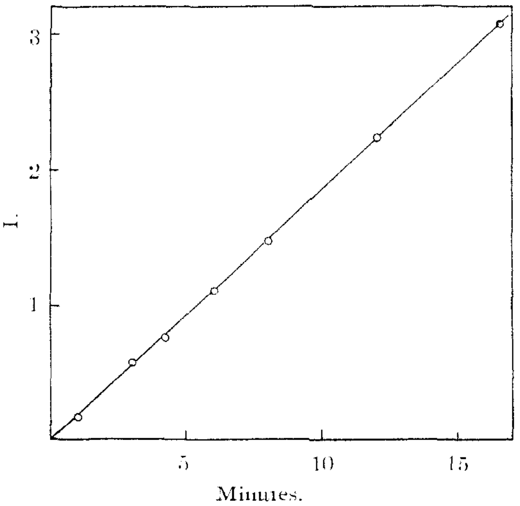


Fig. 2.—Kinetic plot for brominolysis of benzeneboronic acid in "50%" acetic acid at 25.0°: $[Br_2]_0$, 0.00188 M, $[C_6H_5B(OH)_2]_0$, 0.0400 M. Last point corresponds to 22% consumption of bromine; k = 0.00617 l. mole^{-1} sec.$^{-1}$.

Results

Kinetic Order.—Under the conditions used in this investigation the reaction between benzeneboronic acid and bromine is first order with respect to each of the reactants. However, the simple second order rate expression must be modified to take into account the formation of tribromide ion during the course of the reaction. The expression then becomes

$$\frac{dx}{dt} = k[Br_2][RB(OH)_2] = \frac{1}{2}k(c - x)\{\sqrt{(K + b - a + 2x)^2 + 4K(a - x)} - (K + b - a + 2x)\} \quad (1)$$

where a, b and c are the initial concentrations of bromine, bromide ion and substrate, respectively, x is the concentration of product and K is the equilibrium constant for the dissociation of tribromide ion. The integration was performed graphically[7] *using the value 0.020 for K in "50%"* acetic acid.[8]

Figure 1 shows the kinetic plot[9] for a run in which 0.25 M sodium bromide was initially present. Figure 2 shows a similar plot for a typical run in which no sodium bromide was present at the outset. In Table I are collected the data for several individual runs without added bromide ion which were followed to 20-30% consumption of bromine. It is seen that the rate constant increases by about 10% with a fivefold increase from 0.00174 M in initial bromine concentration. This is attributed to volatility loss with increasing bromine concentration. Results obtained in clear and opaque reaction flasks were the same.

Table I

Brominolysis of Benzeneboronic Acid in "50%" Acetic Acid at 25.0°

$[Br_2]_0 \times 10^3$	$[C_6H_5B(OH)_2]_0 \times 10^3$	k(l. mole^{-1} sec.$^{-1}$) $\times 10^3$
1.74	40.0	6.27
1.88	40.0	6.17
3.09	5.00	6.27
3.31	5.00	6.50
8.69	5.00	7.03
8.45	5.00	6.67
5.47	9.94	6.65
6.27	7.95	6.47

Results obtained at an ionic strength of 0.25 in reactions in which sodium bromide was added are summarized in Table II. Here the rate constant does not change significantly throughout the range 7.27 × 10^{-3} – 4.30 × 10^{-4} M in initial free bromine concentration.

When higher initial concentrations of the reactants than those in the experiments reported

---

[6] F. R. Bean and J. R. Johnson, This Journal, **54**, 4415 (1932).

[7] Equation (1) has been integrated (A. E. Bradfield, B. Jones and K. J. P. Orton, *J. Chem. Soc.*, 2810 (1929)) yielding a rather cumbersome expression. We found, however, that graphical integration provided the same results well within the experimental error as evaluation of the integral and in much less time.

[8] W. J. Jones, *ibid.*, **99**, 392 (1911).

[9] $I = \int_0^x \frac{dx}{(c - x)\{\sqrt{(K + b - a + 2x)^2 + 4K(a - x)} - (K + b - a + 2x)\}}$

TABLE II

BROMINOLYSIS OF BENZENEBORONIC ACID IN "50%" ACETIC ACID AT 25.0[a]

$[Br^-]_0$	$[NaClO_4]$	$[Br_2]_0 \times 10^3$	k (l. mole^{-1} sec.$^{-1}$) $\times 10^3$
	0.25	3.76	7.57
	.25	4.75	7.27
0.05	.20	4.32	7.14
.05	.20	3.78	7.06
.15	.10	3.86	6.85
.15	.10	5.37	6.89
.25		7.95	7.16
.25		5.00	6.76

[a] $[C_6H_5B(OH)_2]_0 = 0.0200$ in all experiments.

above were used a decrease in the second order rate constant occurred as the reaction progressed. This is illustrated in Fig. 3. It is evidently due to the protons liberated in the reaction (Equation 1), because the curvature in the plot disappears when a relatively large concentration of perchloric acid is present in the reaction mixture. Furthermore, the effect is not noticeable (due to the law of mass action) if the concentrations of bromine and boronic acid (see Fig. 1) are initially small.

Effect of Solvent.—The rate constant in "30%" acetic acid ($K = 0.023$) is 0.0374 l. mole^{-1} sec.$^{-1}$, and in "20%" acetic acid ($K = 0.028$) it is 0.0738 l. mole^{-1} sec.$^{-1}$. These represent increases of about six- and tenfold, respectively, over the rate in "50%" acid.

Effect of Added Salts.—All of the substances introduced into the reaction mixture altered the rate constant. The results are assembled in Table III. Of the mineral acid salts only perchlorate and nitrate produce the same degree of acceleration. This is regarded as a salt effect. Sulfate and dihydrogen phosphate on the other hand, produce a much greater increase so that their effects must be regarded as catalytic. The two acids investigated both decrease the rate, perchloric acid being much more effective.

TABLE III

EFFECT OF ADDED SUBSTANCES ON THE RATE OF BROMINOLYSIS OF BENZENEBORONIC ACID IN "50%" ACETIC ACID AT 25.0°

Ionic strength = 0.12 except where indicated

Substance added	Concn., moles/l.	$k \times 10^3$, l. mole^{-1} sec.$^{-1}$	k_c, l. mole^{-1} sec.$^{-1}$
None		6.41	
$NaClO_4$	0.12	7.28	
$NaNO_3$	.30 ($\mu = 0.3$)	7.62	
Na_2SO_4	.10 ($\mu = 0.3$)	12.7	
$NaHSO_4$	.30 ($\mu = 0.3$)	5.06	
$HClO_4$	.10 ($\mu = 0.25$)	4.34	
	.12	4.16	
NaH_2PO_4	.12	15.3	0.067
	.06	12.0	.078
CH_3COONa	.12	44.0	.306
	.06	25.0	.292
$ClCH_2COONa$	.12	36.5	.243
	.06	22.8	.258
$Cl_2CHCOONa$	.12	24.0	.139
	.06	14.9	.127
C_6H_5COONa	.12	47.6	.336
	.06	26.4	.318

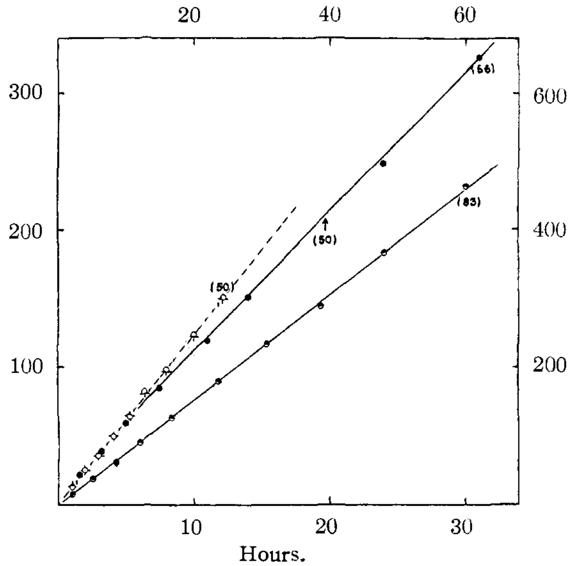


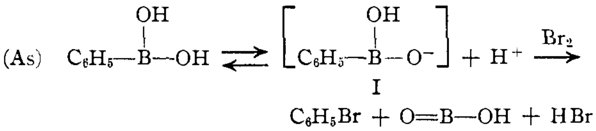
Fig. 3.—Effect of reactant concentrations and added acid on kinetics: 0.02 M $C_6H_5B(OH)_2$, 0.15 M NaBr, 0.25μ; ○, 0.00386 M Br_2; +, 0.00537 M Br_2; ●, 0.01516 M Br_2; ◓, 0.10 M $HClO_4$, 0.00693 M Br_2 (right ordinate, top abscissa). Figures in parentheses indicate per cent. bromine consumed at time indicated.

All of the salts of carboxylic acids have pronounced catalytic effects. In the last column of the table are given the values of the rate constants $k_c = (k - 0.00728/\text{cat. concn.})$ for the base-induced reactions at $\mu = 0.12$. The values obtained at the two concentrations for each show good agreement in most cases. Thus, the rate of brominolysis of benzeneboronic acid in the presence of bases in the range of concentrations studied can be represented by the expression

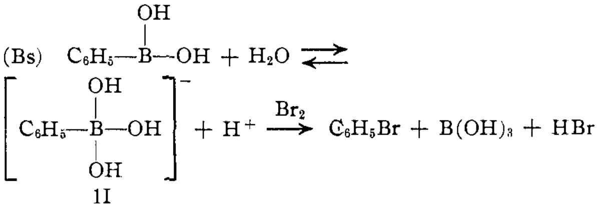
$$\frac{dx}{dt} = \{k_0 + k_c(\text{base})\}\,[\phi B(OH)_2]\,[Br_2]_{\text{free}} \qquad (2)$$

Discussion

The "Solvent Reaction."—The results obtained in this investigation suggest two mechanisms as being particularly attractive.



$$(\text{As})\quad C_6H_5-\overset{\overset{\displaystyle OH}{|}}{B}-OH \rightleftarrows \underset{\text{I}}{\left[C_6H_5-\overset{\overset{\displaystyle OH}{|}}{B}-O^-\right]} + H^+ \xrightarrow{Br_2} C_6H_5Br + O{=}B-OH + HBr$$



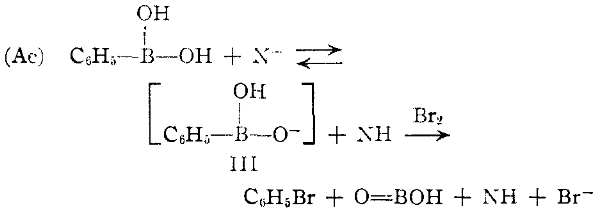
$$(\text{Bs})\quad C_6H_5-\overset{\overset{\displaystyle OH}{|}}{B}-OH + H_2O \rightleftarrows \underset{\text{II}}{\left[C_6H_5-\overset{\overset{\displaystyle OH}{|}}{\underset{\underset{\displaystyle OH}{|}}{B}}-OH\right]^-} + H^+ \xrightarrow{Br_2} C_6H_5Br + B(OH)_3 + HBr$$

In each case the (rapidly attained) equilibrium would be shifted to the right by increased water concentration which leads to a less acidic solvent of higher dielectric constant. The species I or II reacts with bromine in the rate-determining step. Considerable driving force for the reaction is provided by the presence of the negative charge.

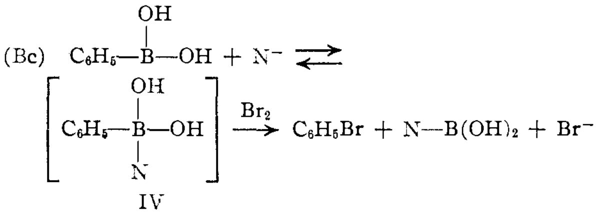
The carbon atom attached to the boron has a greater electron density *via* the inductive effect than it has in the neutral boronic acid molecule and is therefore more susceptible to electrophilic attack.

The effect of acids in decreasing the reaction rate is to decrease the concentration of I or II by shifting the equilibrium to the left.

The Base-Induced Reaction.—Here again two mechanisms in which the role of the solvent in As and Bs is played by a nucleophilic species N^- may be considered.



$$(Ac)\quad C_6H_5-\overset{OH}{B}-OH + N^- \rightleftarrows \underset{III}{\left[C_6H_5-\overset{OH}{B}-O^-\right]} + NH \xrightarrow{Br_2} C_6H_5Br + O{=}BOH + NH + Br^-$$



$$(Bc)\quad C_6H_5-\overset{OH}{B}-OH + N^- \rightleftarrows \underset{IV}{\left[C_6H_5-\overset{OH}{\underset{N}{B}}-OH\right]} \xrightarrow{Br_2} C_6H_5Br + N-B(OH)_2 + Br^-$$

A further possibility would have N^- functioning in displacing the equilibrium of Bs toward the right by removal of protons (C_c).

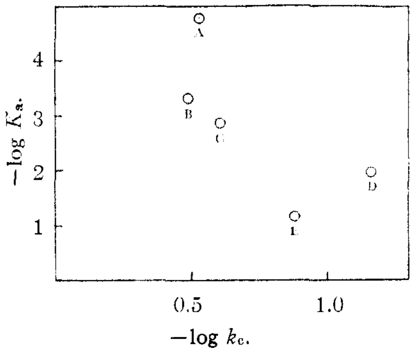


Fig. 4.—Test of Brönsted relation for the base catalysis: A, CH_3OO^-; B, $C_6H_5COO^-$; C, $ClCH_2COO^-$; D, $H_2PO_4^-$; E, Cl_2CHCOO^-.

All of these mechanisms are consistent with the qualitative observations. However, a distinction between Ac and Bc can be made on the basis of the quantitative results. In Ac the base engages in a proton transfer reaction, but in Bc it engages in a nucleophilic attack on boron. Now, relative basicity as measured by affinity for protons appears to be independent of the solvent provided the bases under comparison are of the same charge type.[10,11] Therefore, if mechanism Ac is the only one being followed a plot of $\log k_c$ *vs.* $\log k_b$ ($-\log k_a$) for bases should yield a straight line[10,11] (the Brönsted catalysis law). Figure 4 shows that this is not the case for the singly charged bases.[12]

On the other hand, basicity as measured by proton affinity bears little relationship to nucleophilic activity as measured by efficiency in displacement reactions at carbon atoms.[13,14,15] Thus we are led to the tentative conclusion that mechanism Bc and by analogy, Bw are to be preferred over the others.[16]

The observations made in this investigation suggest a reasonable interpretation for one interesting aspect of the chemistry of the aromatic boronic acids. Most of the electrophilic displacement reactions which have been attempted occur with replacement of the boronic acid group. These include reactions with halogens,[4,17] hydrogen peroxide,[17] water[17] and metallic salts.[17,18] On the other hand, nitration of benzeneboronic acid can be carried out with the predominant formation of either ortho- or meta-nitrobenzeneboronic acid (and little nitrobenzene) depending on the reaction conditions.[17,19] If electrophilic replacement of the boronic acid group in general requires intermediates such as II and IV or their equivalents this latter observation becomes reasonable. The reaction is necessarily carried out in a strongly acid medium in which the concentrations of the intermediates would be extremely small. This would decrease the rate of replacement of the boronic acid group so that it might eventually fall below the rate of replacement of nuclear hydrogen. Other factors being the same, the rate of replacement of hydrogen should be independent of acid concentration because the cleavage of the carbon–hydrogen bond is not a part of the rate-determining step in nitration.[20]

Further studies on the mechanism of the reaction are being made.

DURHAM, NEW HAMPSHIRE RECEIVED FEBRUARY 12, 1951

---

(10) R. P. Bell, "Acid-Base Catalysis," Oxford University Press, London, 1941, pp. 107 ff.

(11) Exceptions to this statement are discussed by L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, pp. 222 ff.

(12) Further evidence is provided by our semi-quantitative observation that sodium fluoride is probably at least a hundred times more effective as a catalyst than sodium acetate although the latter provides a stronger base.

(13) J. N. Brönsted, M. Kilpatrick and M. Kilpatrick, THIS JOURNAL, **51**, 428 (1929).

(14) A. G. Ogston and E. R. Holiday, *Trans. Faraday Soc.*, **44**, 45 (1948).

(15) P. D. Bartlett and G. Small, Jr., THIS JOURNAL, **72**, 4867 (1950).

(16) The possibility that reactions involving species such as $[RB(N)_2OH]^-$ and even $[RB(N)_3]^-$ may occur at higher base concentrations is not excluded.

(17) A. D. Ainley and F. Challenger, *J. Chem. Soc.*, 2171 (1930).

(18) A. Michaelis and P. Becker, *Ber.*, **15**, 180 (1882).

(19) W. Seaman and J. R. Johnson, THIS JOURNAL, **53**, 711 (1931).

(20) L. Melander, *Acta Chem. Scand.*, **3**, 95 (1949); *Nature*, **163**, 599 (1949); *Arkiv. Kemi*, **2**, 213 (1950).