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THE MECHANISM OF THE BROMINATIVE DECARBOXYLATION
OF o- AND p-HYDROXYBENZOIC ACIDS

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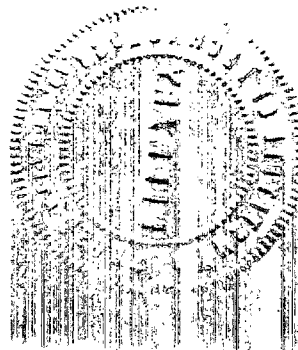
A THESIS

Presented to
the Faculty of the Graduate Division
Georgia Institute of Technology

In Partial Fulfillment
of the Requirements for the Degree
Doctor of Philosophy in the School
of Chemistry

by
Ulysses Virgil Henderson, Jr.

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THE MECHANISM OF THE BROMINATIVE DECARBOXYLATION
OF o- AND p-HYDROXYBENZOIC ACIDS

Approved:

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April 22, 1954

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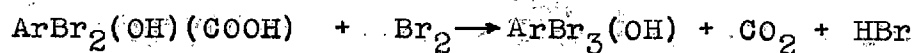
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SUMMARY

When o- and p-hydroxybenzoic acids are treated with bromine in glacial acetic acid, the products of the reaction are the mono- or dibromohydroxybenzoic acid, depending on the amount of bromine used. If the dibromo-acid is reacted with bromine in aqueous acetic acid, carbon dioxide is eliminated and tribromophenol is formed.



This is apparently an electrophilic displacement of the carboxyl group by bromine.

Although electrophilic displacements of aromatic hydrogen atoms have been studied extensively, comparatively few reactions involving the displacement of groups or atoms other than hydrogen have been studied. The broad purpose of this study therefore, was to add to the general fund of information relative to such electrophilic displacements. Specifically, the study was directed toward an elucidation of the mechanism of brominative decarboxylation by a study of the kinetics of the reactions.

After it was shown that at constant hydrogen ion and bromide ion concentrations the rate of the reaction was first order in stoichiometric bromine and first order

in substrate, the kinetics were studied in detail by measuring the effect on the rate of varying hydrogen ion and bromide ion concentrations. The rate was measured by stopping portions of the reaction mixture with potassium iodide at measured time intervals and determining the concentration of unused bromine by titration of the liberated iodine with standard sodium thiosulfate solution.

From the data collected, apparent integrated bimolecular rate constants were calculated from the equation

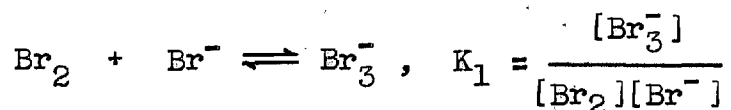
$$k_{app} = \frac{2.303}{(b-a)t} \log \frac{a(b-x)}{b(a-x)}$$

in which b is the initial stoichiometric bromine concentration, a is the initial stoichiometric concentration of hydroxybenzoic acid, t is the time elapsed between the mixing to the reactants and the stopping of the reaction, and x is the concentration of either reactant consumed or the concentration of product formed at time t.

In order to establish the nature of the reacting species, it was necessary to study how the change of concentration of bromide ion and hydrogen ion affected the integrated rate constants. This was done by preparing aqueous acetic acid solvents containing appropriate amounts of lithium bromide, hydrogen bromide, and perchloric acid. The runs made for this purpose were all carried out at an

ionic strength of 0.3 and at 20.0°. The hydrogen ion concentration was then held constant at 0.300 M while the bromide ion concentration was varied, or vice versa.

Before the data could be considered quantitatively, the equilibrium constant for the reaction



had to be evaluated in the various solvents used. This was done spectrophotometrically and K_1 (in liters/mole) was found to be 93 in 80.0 per cent acetic acid, 85 in 75.0 per cent acetic acid, and 83 in 70.0 per cent acetic acid. In this discussion, a term enclosed in square brackets, [], indicates an actual concentration, while a term enclosed in parentheses, (), indicates a stoichiometric concentration.

The rate of bromination of 2,6-dibromophenol was found to be inversely dependent on the first power of the bromide ion concentration and showed an inverse hydrogen ion dependence between zero and first power. The data is consistent with a mechanism in which bromine molecules react with both ionized and un-ionized phenol, or as expressed in a rate equation,

$$\text{rate} = k[\text{Br}_2][(\text{PhOH}) + \frac{k'(\text{PhOH})}{[\text{H}^+]}]$$

This expression can be related to the apparent rate constant by the expression

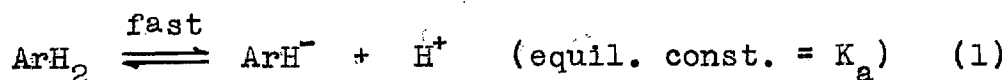
$$\frac{k_{app} K_1 (Br_2)_o [Br^-]_o}{[Br_3^-]_o} = k + \frac{kk'}{[H^+]_o} = F.$$

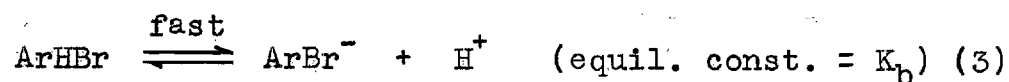
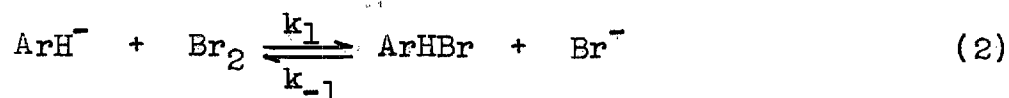
By plotting F vs. $1/[H^+]_o$, a straight line was obtained from which it was determined that $k = 169$ l./mol.min. and $k' = 0.207$ mol./l. The data for such a plot is given in Table I.

Table I

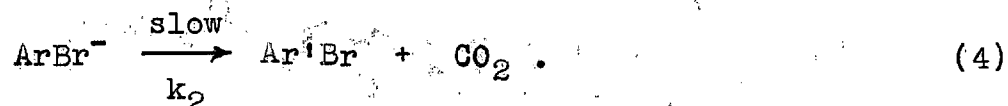
$[H^+]_o$	$(Br_2)_o$	$[Br_3^-]_o$	$[Br^-]_o$	$1/[H^+]_o$	k_{app}	F
0.300	0.00406	0.00392	0.296	3.33	9.61	274
0.300	0.00375	0.00356	0.196	3.33	14.7	282
0.300	0.00371	0.00334	0.0967	3.33	30.8	306
0.200	0.00373	0.00360	0.296	5.00	11.91	340
0.151	0.00798	0.00770	0.292	6.63	13.94	392
0.101	0.00369	0.00356	0.296	9.90	18.11	516

It was found that the rate of bromination of 3,5 di-bromo-2-hydroxybenzoic acid was inversely dependent on the hydrogen and bromide ion concentrations and the dependence in both cases is between first and second power. The data is in agreement with at least two mechanisms. The first is as follows:





and finally



It can be shown that the rate equation for this mechanism is

$$\text{rate} = \frac{k_2 k_1' [\text{Br}_2] (\text{ArH}_2)}{[\text{H}^+]^2 (k_{-1} [\text{Br}^-] + k_2' / [\text{H}^+])} \quad \begin{matrix} (k_1' = k_1 K_a) \\ (k_2' = k_2 K_b) \end{matrix} \quad (5)$$

and this can be related to the apparent second order rate constant by the equation

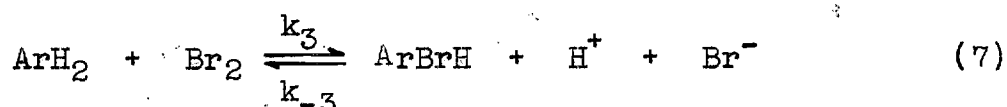
$$\frac{k_{\text{app}} K_1 (\text{Br}_2)_0 [\text{Br}^-]_0 [\text{H}^+]_0}{[\text{Br}_3^-]_0} = \frac{k_1'}{(k_{-1}/k_2' [\text{Br}^-]_0 [\text{H}^+]_0 + 1)} = \Psi \quad (6)$$

If $1/\Psi$ is plotted vs. $[\text{Br}^-][\text{H}^+]$, a straight line is obtained from which it can be determined that $k_1' = 0.42 \text{ min}^{-1}$ and the reversibility ratio, $k_{-1}/k_2' = 34.6 \text{ l.}^2/\text{mol.}^2$. The data for such a plot is given in Table II. The mean deviation of $1/\Psi$ is ± 0.27 .

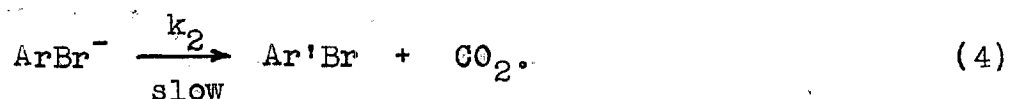
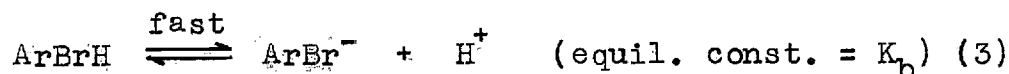
Table II

k_{app}	$[H^+]_0$	$[Br^-]_0$	$[Br_3^-]_0$	$(Br_2)_0$	Ψ
0.0142	0.300	0.292	0.008368	0.008368	0.1098
0.0264	0.300	0.193	0.007488	0.007954	0.138
0.0825	0.300	0.093	0.007075	0.007973	0.2205
0.215	0.300	0.044	0.006300	0.007996	0.306
0.830	0.300	0.0070	0.002968	0.007932	0.396
1.65	0.300	0.000	0.000	0.007942	0.495
0.0257	0.200	0.292	0.007656	0.007964	0.1327
0.0689	0.101	0.292	0.007593	0.007898	0.1801

The second mechanism in agreement with the data for dibromosalicylic acid shows that although the course of bromine and bromide ion is probably unequivocal, other courses for the hydrogen ion are possible. Thus, if the reaction were initiated on the un-ionized acid:



and then as previously shown



The rate equation for this mechanism is

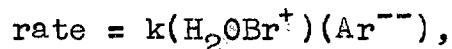
$$\text{rate} = \frac{k_2' k_3 (ArH_2) [Br_2]}{k_{-3} [H^+]^2 [Br^-] + k_2'} \quad (k_2' = k_2 K_b) \quad (8)$$

and this can be related to the apparent second order rate constant by the expression

$$\frac{k_{app}K_1(Br_2)_0[Br^-]_0}{[Br^-]_0} = \frac{1}{(k_{-3}/k_2'k_3)[H^+]_0^2[Br^-]_0 + 1/k_3} = \frac{1}{Q} \quad (9)$$

A plot of $1/Q$ vs. $[H^+]^2 Br^-$ is a straight line from which it can be determined that $k_3 = 1.7$ l./mol.min. and the reversibility ratio $k_{-3}/k_2' = 143$ l.³/mol.³. The data for such a plot can be obtained from Table II. The mean deviation of $1/Q$ is ± 0.11 .

It was found that the rate of bromination of 3,5-dibromo-4-hydroxybenzoic acid was inversely dependent on the second power of both hydrogen and bromide ion concentrations. This is in accord with a mechanism in which HOBr reacts with the mono-ionized form of the acid or a mechanism in which H_2OBr^+ or Br^+ reacts with the di-ionized form of the acid, i.e.,



or



Theoretical considerations seem to indicate that a reaction with H_2OBr^+ is the more probable of these three.

It can also be shown that the mechanism expressed by equations 1-4 for dibromosalicylic acid and culminating in equation 6 holds equally well for dibromo-p-hydroxybenzoic

acid. If $1/\Psi$ is plotted vs. $[\text{Br}^-][\text{H}^+]$, a straight line is obtained from which it can be determined that $k_1' = 6.3 \text{ min}^{-1}$ and the reversibility ratio, $k_{-1}/k_2' = 600 \text{ l.}^2/\text{mol.}^2$. The data for this plot are given in Table III. The mean deviation of $1/\Psi$ is ± 0.13 .

Table III

k_{app}	$[\text{H}^+]_0$	$[\text{Br}^-]_0$	$[\text{Br}_3^-]_0$	$(\text{Br}_2)_0$	Ψ
0.0135	0.300	0.292	0.00777	0.00806	0.114
0.0297	0.300	0.192	0.00780	0.00824	0.168
0.111	0.300	0.093	0.00732	0.00817	0.321
0.456	0.300	0.043	0.00653	0.00813	0.682
(15)	0.300	0.000	0.00000	0.00819	4.5
0.0313	0.200	0.292	0.00795	0.00825	0.175
0.120	0.101	0.292	0.00783	0.00812	0.341
0.414	0.054	0.292	0.00786	0.00816	0.666

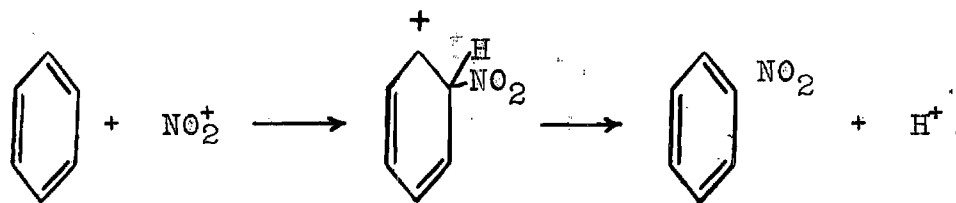
The data available cannot distinguish between mechanisms of the two general types given for the para compound. Because of the uncertainty in $1/\Psi$ especially near zero bromide and hydrogen ion concentrations, the value of k_{-1}/k_2' is only known to be somewhere between 300 and infinity. It can be shown that if the first step (equation 2) is at equilibrium (k_{-1}/k_2' is a very large number), the two mechanisms will reduce to the same rate equation. If k_{app} were known more accurately at low values of $[\text{H}^+][\text{Br}^-]$, it might be possible to distinguish between the two mechanisms.

Since the rate of bromination of both 2,4- and 2,6-dibromophenol was so much faster than that of the hydroxybenzoic acids, a mechanism involving a fast displacement of the carboxyl group by a proton followed by a rate controlling bromination of the resulting phenol was proven to be unimportant in the brominative decarboxylation of the hydroxybenzoic acids. The bromine dependence of the brominations likewise disproved a rate controlling displacement of the carboxyl group by a proton.

CHAPTER I

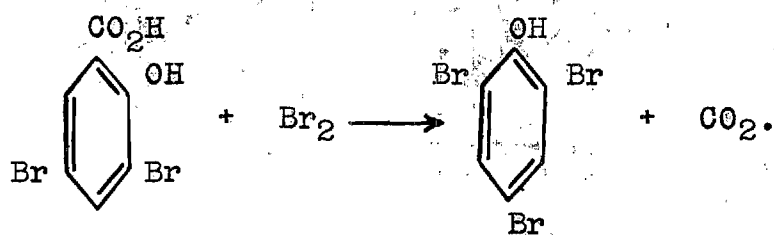
INTRODUCTION

Ordinary aromatic substitution is a reaction wherein an electrophilic reagent displaces a nuclear hydrogen atom from an aromatic compound, e.g.,



However, aromatic electrophilic substitutions are not limited to the displacement of hydrogen atoms. When the compound is suitably activated, other atoms or groups may be displaced by an electrophile. The particular reactions of interest in this study are of the latter type, in which a carboxyl group is displaced by halogen.

When ortho- or para-hydroxybenzoic acid is brominated with excess bromine in glacial acetic acid, the product formed in almost quantitative yield is the appropriate dibromohydroxybenzoic acid. When this dibromo-acid is treated with bromine in a more polar solvent (containing a strong acid), tribromophenol is formed quantitatively, e.g.,



It is the purpose of this study to measure the kinetics of this reaction under various conditions and to set forth a mechanism (or mechanisms) for the reaction consistent with the observed data.

CHAPTER II

SURVEY OF RELATED WORK

Halogenation of hydroxybenzoic acids.--While studying the reactions of salicylic acid, Cahours (1) noted that when salicylic acid was treated with bromine drop by drop in the absence of a solvent, he could isolate mono-, di-, and tri-bromosalicylic acid and in addition he noted some bromophenols. The reaction conditions were obviously rather vigorous.

Similarly, Peltzer (2) noted that when free p-hydroxybenzoic acid was subjected to the direct action of chlorine, especially in the presence of sunlight or at high temperature, a mixture of chlorinated products were obtained. He prepared chloro-p-hydroxybenzoic acid by the action of chlorine on the dry silver salt in the absence of heat or light. He also studied the action of iodine on both p-hydroxybenzoic acid and its methyl ether, anisic acid. He reported that when the former was treated with iodine and iodic acid in boiling water, both 3-iodo- and 3,5-diiodo-4-hydroxybenzoic were formed; however, when anisic acid was treated with the same reagents at 145-150° for six hours, only 3-iodo-4-

-
1. M. A. Cahours, Ann. chim. phys., (3), 13, 87 (1845).
 2. R. Peltzer, Ann., 146, 284 (1868).

methoxybenzoic acid resulted. Peltzer did not report that he observed any decarboxylation of these compounds by iodine.

On the other hand, when Weselsky (3) studied the iodination of the hydroxybenzoic acids in 90 per cent alcohol solutions, he did observe decarboxylation. Using iodine and mercuric oxide as the reagent he obtained both the mono- and diiodo-acids from the ortho- and para-hydroxybenzoic acids. They were purified and separated by using boiling water and in this process both diiodo- and triiodophenol were found.

Hubner and Heinzerling (4) observed that the bromohydroxybenzoic acid obtained from bromonitrobenzoic acid by reduction of the nitro group, followed by diazotization and hydrolysis, was the same as that formed by bromination of salicylic acid in carbon disulfide. They therefore deduced that bromosalicylic acid was a derivative of phenol and Hubner and Brenken (5) were able to prove this by a dry distillation of bromosalicylic acid, the solid product of which was monobromophenol. This work was later verified and extended by Lellmann and Grothmann (6).

In the course of studying the unusual products that

3. P. Weselsky, Ann., 174, 99 (1874).

4. H. Hubner and C. Heinzerling, Zeitschrift für Chemie, 7, 709 (1871).

5. H. Hubner and O. Brenken, Ber., 6, 170 (1873).

6. E. Lellmann and R. Grothmann, Ber., 17, 2724 (1884).

could be obtained by the bromination of phenols, Benedikt (7) observed that when either *p*-hydroxybenzoic acid or salicylic acid was treated in aqueous solution with an excess of bromine water, that tribromophenol bromide was formed in good yield. He studied the properties of this compound extensively and noted that the fourth bromine atom is very labile and the compound will brominate phenol or aniline.¹

The bromination and chlorination of the three isomeric hydroxybenzoic acids was studied extensively by Zinke and Walbaum (8) in an attempt to find isomeric products that could possibly prove the existence of Kekule type isomers of benzene. Most of their work was done in glacial acetic acid

7. R. Benedikt, Ann., 199, 127 (1879).

8. T. Zinke and H. Walbaum, Ann., 261, 208 (1891).

¹The exact structure of tribromophenol bromide has been the subject of much controversy and experimentation since Benedikt's discovery of it. Benedikt postulated that the substance was a hypohalite and has been supported in this view by W. M. Lauer (J. Am. Chem. Soc., 48, 442 (1926)), I. Suknevich and S. Budnitzkii (J. prakt. Chem., (N.F.) 138, 18 (1933)), and G. Wittig and F. Vidal (Ber., 81, 368 (1948)). On the other hand, many workers have favored a cyclohexadienone structure, e.g., J. Thiele and H. Eichwede (ibid., 33, 673 (1900)), J. H. Kastle and R. Speyer (Am. Chem. J., 27, 40 (1902)), J. H. Kastle and A. S. Loevenhart (ibid., 27, 32 (1902)), and J. H. Kastle and J. W. Gilbert (ibid., 27, 43 (1902)). Several workers feel that the two structures are only labile tautomers, e.g., L. C. Raiford and A. L. LeRosen (J. Am. Chem. Soc., 66, 2080 (1944); ibid., 68, 397 (1946)), and C. H. R. Elston, A. T. Peters, and F. M. Rowe (J. Chem. Soc., 1948, 367).

Tribromophenol bromide has occasionally been involved in procedures for the determination of phenols and this problem has been discussed at length by S. J. Lloyd (J. Am. Chem. Soc., 27, 7 (1905)), S. C. J. Olivier (Rec. trav. chim., 28, 354 (1909)), and I. M. Kolthoff (Pharm. Weekblad, 69, 1147 (1932)).

and under such a wide range of conditions that they were able to obtain products ranging from the monohalo-acid to a penta- and hexahaloketo-acid. They did not, however, observe a spontaneous displacement of the carboxyl group by halogen. They did carry out thermal eliminations of the carboxyl group of the alkali salts to obtain certain desired phenols.

Schunck and Marchlewski (9), while studying the bromination of the natural product datiscetin, noticed a cleavage in which salicylic acid or bromosalicylic acid was formed. Since the bromination of datiscetin also gave bromanil, they reasoned that perhaps salicylic acid might be the intermediate in the formation of bromanil. This was confirmed by suspending salicylic acid in acetic acid and adding excess bromine. A vigorous reaction ensued in which carbon dioxide was evolved. If the mixture was heated to boiling for a short time, bromanil could be isolated. Ultimately they reported that anisic and *p*-hydroxybenzoic acids could also be reacted with bromine to form bromanil with the same facility. The term "same facility" is regarded skeptically because due to the vigorous nature of the reaction conditions, it is doubtful that an accurate comparison of reactivity could be made.

9. E. Schunck and L. Marchlewski, Ann., 278, 349 (1894).

Lössner (10) chlorinated both salicylic and *p*-hydroxybenzoic acid by warming them with antimony pentachloride in the absence of a solvent. He did not report that he observed any decarboxylation; however, his research was directed solely toward the use of antimony pentachloride as a reagent for organic chlorination.

The classical method of preparing dichlorosalicylic acid was for many years the method reported first by Smith (11). He used glacial acetic acid as the solvent for the chlorination as well as for the purification of the product. Lellmann and Grothmann (6) likewise found glacial acetic acid to be a suitable solvent for the bromination of salicylic acid. Earle and Jackson (12) have published procedures for the preparation of both dichloro- and dibromosalicylic acids in glacial acetic acid. Their procedures are essentially revisions of the methods of Smith (11) and Lellmann and Grothmann (6). Robertson (13) has given a similar preparation for 3,5-dibromosalicylic acid and in addition has also reported a preparation in the same solvent using iodine as a catalyst. The advantages of using glacial acetic acid as a solvent are that decarboxylation is minimized and the monohalo acid is soluble in the solvent.

10. C. W. Lössner, J. prakt. Chem., (2), 13, 418 (1876).

11. E. J. Smith, Ber., 11, 1225 (1878).

12. R. B. Earle and H. L. Jackson, J. Am. Chem. Soc., 28, 104 (1906).

13. W. Robertson, J. Chem. Soc., 81, 1475 (1902).

Pope and Wood (14) prepared 3,5-dibromo-4-hydroxybenzoic acid by dissolving p-hydroxybenzoic acid in 75 per cent sulfuric acid cooled to zero degrees and then adding bromine in a cold glacial acetic acid solution. The preparation requires one week and results in the formation of a small amount of tribromophenol. Pope and Wood studied the thermal elimination of carbon dioxide from dibromo-p-hydroxybenzoic acid in the presence of acids, water, and base; but they did not study the displacement of the carboxyl group by halogen.

Brenans and Girod (15) have prepared bromiodophenols and chloriodophenols from 5-bromo- and 3,5-dibromosalicylic acids and the chlorosalicylic acids by the action of iodine in alkaline solutions.

Fuming sulfuric acid was used as solvent in the preparation of a number of halogenated derivatives of salicylic acid by Farinholt, Stuart, and Twiss (16). In this medium there is apparently no decarboxylation. They were able to prepare a tribromo- and a tetrabromosalicylic acid but could not chlorinate any higher than trichlorosalicylic acid. They were also able to prepare monobromotrichlorosalicylic acid.

14. F. G. Pope and A. S. Wood, J. Chem. Soc., 101, 1823 (1912).

15. P. Brenans and C. Girod, Compt. rend., 186, 1128 (1928).

16. L. H. Farinholt, A. P. Stuart and D. Twiss, J. Am. Chem. Soc., 62, 1237 (1940).

When 3,4,5,6-tetrabromosalicylic acid was reacted with bromine in 30 per cent acetic acid at 60°, 2,3,4,5,6-pentabromophenol was formed. In a like manner 2,3,4,6-tetrabromophenol was formed from tribromosalicylic acid. The chlorosalicylic acids behave analogously.

Kolthoff (17) has studied the bromination of salicylic acid as a method of analysis for salicylic acid. He observed that in water, salicylic acid or salicylate ion reacts with bromine water to form first dibromosalicylic acid, then decarboxylation takes place with substitution of a third bromine atom to form tribromophenol which finally takes up a fourth bromine atom to yield tribromophenol bromide. The carbon dioxide is liberated quantitatively and may be absorbed in barium hydroxide for volumetric or gravimetric data. This reaction may be employed for the determination of salicylic acid in the presence of other phenols.

17. I. M. Kolthoff, Pharm. Weekblad, 69, 1159 (1932).

Halogenation of aminobenzoic acids.--The decarboxylation of p-aminobenzoic acid by bromine water had been noted as early as 1866 by Beilstein and Geitner (18). They observed that tribromoaniline was formed in the reaction as well as 3,5-dibromo-4-aminobenzoic acid. Similarly, when Bogert and Hand (19) prepared 3,5-dibromo-2-aminobenzoic acid by dissolving anthranilic acid in a large volume of dilute hydrochloric acid and adding slightly less than an equivalent amount of bromide-bromate reagent, they observed that tribromoaniline was one of the products.

When Wheeler and Liddle (20) tried to prepare 3-iodo-4-aminobenzoic acid by reacting an aqueous solution of the potassium salt of p-aminobenzoic acid with two atomic proportions of iodine, they obtained only p-iodoaniline. They were able to prepare the 3-iodo-4-aminobenzoic acid by passing one molar proportion of ICl vapor into a solution of p-aminobenzoic acid in a small volume of cold hydrochloric acid. If, however, the p-aminobenzoic acid was dissolved in a large amount of dilute hydrochloric acid and one molar proportion of ICl was added and the solution warmed or concentrated, they obtained 3,5-diiodo-4-aminobenzoic acid. When aqueous solution of the latter was boiled with powdered iodine,

18. F. Beilstein and P. Geitner, Ann., 139, 1 (1866).

19. M. T. Bogert and W. F. Hand, J. Am. Chem. Soc., 25, 935 (1903).

20. H. L. Wheeler and L. M. Liddle, Am. Chem. J., 42, 441 (1909).

no action was apparent, but when the mixture was heated in a closed tube at 165° for five hours, a black precipitate was obtained which was found to be triiodoaniline upon purification.

Analogous results were observed by Wheeler and Johns (21) when they tried to prepare 3,5-diiodo-2-aminobenzoic acid. When they added iodine to an aqueous solution of the potassium salt of 5-iodoanthranilic acid, only triiodoaniline was obtained. They found, however, that the diiodoanthranilic acid could be prepared by iodination of either 5-iodoanthranilic acid or anthranilic acid in a solution of about 5 per cent hydrochloric acid. This procedure cut the yield of triiodoaniline to less than 10 per cent.

Ullmann and Kopetschni (22) similarly discovered that they could prepare 3,5-dibromoanthranilic acid by brominating anthranilic acid in an acidic aqueous solution. They noted that it is soluble in concentrated hydrochloric acid and is transformed into tribromoaniline by bromine water, but that the carboxyl group is not displaced in strong fuming sulfuric acid.

When Wheeler and Oates (23) brominated anthranilic acid in cold glacial acetic acid (near the freezing point of the mixture) they found that two thirds of the product

21. H. L. Wheeler and G. O. Johns, Am. Chem. J., 43, 398 (1910).

22. F. Ullmann and E. Kopetschni, Ber., 44, 425 (1911).

23. A. S. Wheeler and W. M. Oates, J. Am. Chem. Soc., 32, 770 (1910).

consisted of 5-bromoanthranilic acid and one third of the dibromo-acid. When they brominated anthranilic acid in near boiling glacial acetic acid, the results were reversed, giving two thirds of the 3,5-dibromoanthranilic acid. The two acids are readily separated by boiling water in which the dibromoanthranilic acid is nearly insoluble. They report that a little tribromoaniline was found after brominating in cold acetic acid. One would have expected that tribromoaniline would more likely be found after brominating in hot acetic acid.

Sudborough and Lakhumalani (24), contrary to other workers, claim to have observed extensive decarboxylation of aminobenzoic acids in glacial acetic acid. They reported that when a glacial acetic acid solution of 3,5-dibromo-2-aminobenzoic acid was treated with an acetic acid solution of chlorine, a mixture of chlorobromoanilines was obtained. In the same solvent it was observed that 3,5-dibromo-4-aminobenzoic acid yielded 90 per cent tribromoaniline when treated with bromine, but when treated with chlorine no products could be isolated. Since other workers almost unanimously report that decarboxylation is minimized in glacial acetic acid, it seems likely that these workers were not really using anhydrous acetic acid.

24. J. J. Sudborough and J. V. Lakhumalani, J. Chem. Soc., 111, 41 (1917).

A practical application of the decarboxylation reaction has been reported by Wells (25). The Association of Official Agriculture Chemists method for the determination of procaine (and similar compounds) depends on the quantitative bromination in aqueous solvent of the p-aminobenzoic acid formed by hydrolysis to tribromoaniline. With p-aminobenzoic acid, quantitative bromination was obtained when 140 per cent excess of bromine was used and two hours were allowed for the reaction to reach completion. The time required varied inversely with the excess of bromine used.

25. E. H. Wells, J. Asso. Official Agr. Chem., 25, 537 (1942).

Survey of the mechanisms of related halogenation reactions.--

A considerable amount of work on the mechanism and rate of bromination was done by Francis and co-workers (26). He was primarily concerned with using bromination as an analytical device for the estimation of the amount of meta isomer present in mixtures of monosubstituted anilines, phenols, and nitro compounds. A difficulty in the method was the displacement of $-COOH$, $-CHO$, and $-SO_3H$ by bromine when these are present. He found that if the reaction mixture was cooled below $0^\circ C.$, that displacement was halted. The brominations were done in aqueous solutions with enough alcohol added to prevent the precipitation of the dibromocompound, and if carried out at $20^\circ C.$ or above, displacement of the group is quantitative in almost every case.

The determination of the relative rates of bromination of certain ortho-, meta-, and para- isomers of monosubstituted phenols and anilines was undertaken by Francis (27, 28). He found that bromination by free bromine in water is too rapid to study in the case of aniline or

26. A. W. Francis and A. J. Hill, J. Am. Chem. Soc., 46, 2498 (1924).

27. A. W. Francis, A. J. Hill, and J. Johnston, J. Am. Chem. Soc., 47, 2211 (1925).

28. A. W. Francis, J. Am. Chem. Soc., 48, 1631 (1926).

phenol, and if bromide-bromate reagent is used, the rate observed is that of the formation of molecular bromine, which is slower than that of the bromination of aromatic compounds with an amino or hydroxyl group. By means of competition experiments using two different compounds at a time with a limited amount of bromine and a determination of the relative amounts of each product formed, he was able to determine the relative rates of entrance of the first, second, and third bromine atoms into a number of compounds and the relative rates of bromination of the different compounds. His observed ratios are practically independent of concentration and temperature, and probably catalysts, but they vary greatly with changes in the alcohol or acid concentration, since the competitor which is brought nearer to saturation by the change in medium is favored in the competition for bromine. Phenols and anilines should not be compared due to the different effect by acid.

As a result of his studies on the relative rates of bromination, Francis became interested in discovering which species actually constituted the active agent in bromination (29). He found that m-nitrophenol was brominated by free bromine water nearly 1000 times as rapidly as by an equivalent concentration of HOBr from which the concentration of

29. A. W. Francis, J. Am. Chem. Soc., 47, 2340 (1925).

free bromine had been reduced to an extremely small value by silver sulfate. The ratio is approximately that of the equilibrium concentration of free bromine in the solution as calculated from the hydrolysis constant of bromine and the solubilities of silver bromide and silver sulfate. Based on the products formed by the addition of bromine and HOBr to olefins in aqueous solutions and in the presence of salts and strong acids, Francis concluded that the active species is the same in both cases and is probably a positively charged bromine atom.

Soper and Smith (30) have studied the halogenation of phenols in aqueous alkaline media using HOCl as the reagent. For the simple phenols, kinetic data could not distinguish between the reaction being between phenol and hypochlorite ion or between phenoxide ion and HOCl. By studying instead the reaction between salicylic acid and HOCl in different hydroxide ion concentrations, they found that only the velocity constant calculated from a reaction between the doubly ionized salicylic acid and un-ionized HOCl was independent of the hydroxide concentration of the medium.

Soper and Smith have shown that the ratio of the ionization constant of Cl_2 to that of HOCl (Cl^+ being the positive ion in each case) is equal to the ion product

30. F. G. Soper and G. F. Smith, J. Chem. Soc., 1926, 1582.

constant of water divided into the hydrolysis constant of chlorine and is equal to 4.5×10^{10} to one. If Cl^+ is the active agent in each case, Cl_2 would be expected to chlorinate 4.5×10^{10} faster than HOCl . The corresponding values for Br_2 and I_2 and their hypohalous acids are 5×10^5 and 30 to 1 respectively. Thus while HOCl would have little chlorinating effect in comparison with free chlorine, we would expect the effects of I_2 and HIO to be of the same order of magnitude.

In order to test the theory of halogenation by positive halogen ion, the chlorination of phenols has been studied by Soper and Smith, using HOCl and Cl_2 as chlorinating agents. The much greater stability of HOCl (as compared with HOBr and HIO) has allowed a sharp differentiation between the mechanisms of substitution of phenols by these two chlorinating agents. In the chlorination of phenols and phenolic ethers by HOCl , the presence of a small amount of HCl has a great effect (increase) on the rate of interaction of HOCl and a phenol. The rate is independent of the concentration and nature of the phenol and is dependent on the concentration of HOCl and the square of the concentration of HCl , thus it is obvious that the rate measured is that of the formation of chlorine from HCl and HOCl , the chlorine reacting immediately with the phenolic substance present. HOCl does not react with

phenolic ethers, since they do not form a phenoxide ion, but Cl_2 reacts easily with both phenols and ethers. It is probable, however, that the interaction of Cl_2 and phenol involves both the highly reactive phenoxide ion and the un-ionized phenol.

It is the opinion of Soper and Smith that the theory of the reactivity of phenoxide ion with un-ionized HOCl that they have demonstrated for the interaction of HOCl and phenols also applies to the similar reactions involving HOBr and HOI . They found, for example, that the rate of iodination is decreased in strong alkaline solution, which is explained by the great decrease in the concentration of un-ionized HOI .

Bradfield, Jones and Orton (31) found that in 50 per cent aqueous acetic acid (by volume), HBr either added or formed during the course of the reaction, causes marked retardation in the bromination of phenol ethers with bromine. Assuming that only the "free" bromine is active in bromination, the appropriate kinetic expression applied to the bromination of phenolic ethers in experiments in which 5 or 10 molecular proportions of HBr were present at the start, gave bimolecular constants (the expression was first power in bromine and in ether) which showed a slight downward drift as the reaction progressed. With smaller original amounts of HBr ,

31. A. E. Bradfield, B. Jones, and K. J. Orton, J. Chem. Soc., 1929, 2810.

the drift in constants becomes more marked. The appropriate expression mentioned above is one that accounts for that part of the bromine originally added which is converted to Br_3^- by the bromide ion present or being formed. The addition of NaBr has the same effect as HBr, however the addition of beta-naphthalene sulfonic acid produced no change in rate.

Although the work of Bradfield and co-workers on the phenol ethers indicated that Br_3^- was a very poor brominating agent, Alexander (32) found evidence that showed otherwise. By laying down p-hexadecylphenol as a monomolecular film on water, he was able to measure the rate of bromination by dilute bromine water. The kinetics of the reaction were pseudo-unimolecular. The addition of potassium bromide caused a pronounced increase in velocity, which increase he assumed must be ascribed to the formation of tribromide ions. From the rates of reaction with bromine water and varying amounts of bromide ion, and with HOBr, he has calculated that Br_3^- is four times as effective as the HOBr molecule, and in a like manner that Cl_3^- is 10^3 more active than HOCl.

In contrast, Bradfield and Jones (33) found that in the reaction between phenolic ethers and chlorine in 99 per cent acetic acid, the reaction is bimolecular and the amount of HCl present has little effect. It does not appear, therefore, that in this medium the trichloride ion plays any

32. A. E. Alexander, J. Chem. Soc., 1938, 729.

33. A. E. Bradfield and B. Jones, J. Chem. Soc., 1928, 1006.

significant part in the reaction, or indeed, that any mechanism other than a direct interaction between chlorine and ether comes into play. The rate is, however, very sensitive to the amount of water in the medium, the speed of chlorination of p-chloroanisole being increased by about 66 per cent as the amount of water in the medium is increased from one per cent to two per cent.

In this same paper, Bradfield and Jones report that in the investigation of the rate of chlorination of several ethers of p-hydroxybenzoic acid, evidence was sought for the possible displacement of the carboxyl group by chlorine in the ratio of 1:1, there was some slight evidence (which was within the experimental error, however) for this displacement. When the proportion of ether to chlorine is made 3:1 or greater, the disturbance in the rate constants disappears, indicating that there is no significant displacement of the carboxyl group.

Shilov and Kanyaev (34) measured the rate of bromination of anisole m-sulfonic acid by HOBr (it should be recalled that Soper and Smith (30) claimed that phenol ethers could not be chlorinated with HOCl). Upon varying the hydrogen ion concentration, they found that the rate of the reaction was proportional to the concentration of the anisole times the

34. E. Shilov and N. Kanyaev, Compt. rend. acad. sci. URSS, 24, 890 (1939); Chemical Abstracts, 34, 4062 (1940).

concentration of the HOBr times the concentration of the hydrogen ion. They suggested that the product of the HOBr and the hydrogen ion concentrations was a measure of the formation of Br^+ or H_2OBr^+ . They have assigned the various brominating species the following relative rate constants: HOBr, 0.12; Br_2 , 80; BrCl , 43,000; and Br^+ , 110,000.

Robertson, de la Mare, and Johnston (35) report that in the bromination of acetanilide, aceto-p-toluidide, mesitylene, anisole, and p-tolyl methyl ether in acetic acid at room temperature in the concentration region of M/40, third order kinetics are observed. This reaction, which has a low heat of activation, changes over to a bimolecular reaction at lower concentrations. Addition of water to the solvent favors a bimolecular change, and addition of chloroform or carbon tetrachloride causes the initiation of chain reactions, with reaction order greater than third. They believe that the termolecular reactions in concentrated solutions must be due to the formation of $(\text{Br}_2)_2$ in pure acetic acid.

Wilson and Soper (36) have reinvestigated the bromination of o-nitroanisole with bromine water and HOBr, and in addition have studied the bromination of benzene under the same conditions. They have observed that precise work on

35. P. W. Robertson, P. B. D. de la Mare, and W. T. G. Johnston, J. Chem. Soc., 1943, 276.

36. W. J. Wilson and F. G. Soper, J. Chem. Soc., 1949, 3376.

the reactivity of aqueous HOBr is handicapped by its decomposition into HBrO_3 and HBr , the latter reacting with HOBr to produce Br_2 , and that this instability necessitates a careful selection of reactions so that the speed of the reaction, while not too rapid for examination, is sufficiently fast to avoid complications by appreciable decomposition of the HOBr. Reactions of HOBr with o-nitroanisole and benzene have been found satisfactory in this respect.

In the study of the bromination of the o-nitroanisole with bromine water and varying the concentrations of bromide ion and hydrogen ion, they found that the rate was not effected by hydrogen ion but was retarded by increasing concentrations of bromide ion. Rate constants were found to be constant when they were calculated on the basis of the free bromine present (that not complexed into Br_3^- by the added Br^-). Hydrolysis of the Br_2 to HOBr was reduced to negligible extent in these studies by the addition of hydrogen ion as well as by the bromide ion present. Since hydrogen ion was found to have no effect on the rate of bromination by bromine water, the bromination cannot be due to the presence of equilibrium HOBr per se. The fact that the specific rate constants, after correction for the presence of inactive Br_3^- , are unaffected by a 100 fold change in the concentration of bromide ion, further indicates that the activity of molecular bromine is not

due to the presence of small amounts of Br^+ formed by the ionization of Br_2 , for this ionization would be greatly affected by the increase in bromide ion concentration.

A comparison of the effect of the hydrogen ion concentration on the rate of bromination by HOBr and bromine was demonstrated by the fact that when the hydrogen ion concentration is increased six fold, the specific rate of bromination by HOBr increases six fold also while the specific rate of bromination by bromine water is unchanged over a thirty fold change of hydrogen ion concentration. This effect is observed with both o-nitroanisole and with benzene. In neither case can the effect be attributed to change in reactivity of the substance undergoing substitution, and must, therefore, be associated with the ionization of HOBr , causing an increased concentration of active Br^+ , or more likely, H_2OBr^+ .

Wilson and Soper have also observed that weak acid buffers catalyze bromination of o-nitroanisole and benzene by HOBr , but at a pH of 2.74 they exhibit no effect on the bromination of these compounds by bromine water. Evidently, inappreciable amounts of acyl hypobromites are formed in such acid solutions.

De la Mare, Hughes, and Vernon (37) have found chlorination by HOCl in the presence of sulfuric or perchloric

37. P. B. D. de la Mare, E. D. Hughes, and C. A. Vernon, Research (London), 3, 192 (1950).

acids to be analogous to nitration in some cases. For adequate concentrations of fairly reactive compounds, the kinetically fundamental process is the heterolysis of the Cl-O bond. Under conditions which effect a rapid enough removal of the chlorine cation (Cl^+), the rate is independent of the concentration of the aromatic substance. With less reactive compounds, or insufficient concentrations of compounds, so that the chlorine cation is allowed to approach a stationary concentration, the rate becomes dependent on the concentration of the aromatic compound.

In studying the bromination of phenolic ethers in 75 per cent acetic acid, Bradfield, Davies, and Long (38) have found kinetic evidence that the rate expression can best be represented by the sum of two terms,

$$dx/dt = k_p(E)(\text{Br}_2) + k_t(E)(\text{Br}_2)^2.$$

They found that HOBr is a powerful brominating agent for ethers in this medium but they were unable to reconcile bromination by HOBr as causing the termolecular term.

They feel that a contribution by $(\text{Br}_2)_2$, as suggested by Robertson and co-workers (35), satisfactorily explains this term.

Berliner (39) has studied the kinetics of the iodination of aniline with iodine in aqueous solution. He has

38. A. E. Bradfield, G. I. Davies, and E. Long, J. Chem. Soc., 1949, 1389.

39. E. Berliner, J. Am. Chem. Soc., 72, 4003 (1950).

found that the rate is independent of the hydrogen ion concentration for pH's greater than 5.4 and inversely proportional to the square of the iodide ion concentration, and to show general base catalysis. The kinetics of the reaction are not compatible with iodination of aniline by I^- , I_3^- , I_2 , HOI , or OI^- . The kinetics are compatible with iodination of aniline by the iodine cation. However, the same kinetics are also in agreement with a reaction between hypiodous acid and the anilinium ion. Kinetically, the two reactions are indistinguishable; however, the latter can be excluded on chemical grounds, since the anilinium ion is a strong meta director and a minimum yield of 64 per cent p-iodoaniline was isolated from the reaction mixture. Furthermore, at the pH at which the concentration of anilinium ion becomes appreciable the rate falls off and would presumably reach zero when all free aniline is converted to the anilinium ion. Berliner feels that it is highly unlikely that these results can be generalized to include Br^+ or Cl^+ as the active agents for Br_2 and Cl_2 in the absence of a catalyst.

Berliner (40) has also studied the kinetics of the iodination of phenol under the same conditions that he studied the iodination of aniline (39) and has attempted to compare the two reactions theoretically. It seems probable from his results that the phenoxide ion is the reactive

40. E. Berliner, J. Am. Chem. Soc., 73, 4307 (1951).

species and the kinetics are compatible with a general acid catalyzed iodination by HOI on the phenoxide ion. From this consideration, he has reevaluated his ideas on the iodination of aniline and has decided that perhaps the correct kinetic expression involves aniline, hypiodous acid, and a proton (in a general acid catalyzed reaction) and in which the term $(\text{HOI})(\text{H}^+)$ is merely the hydrated iodine cation. The function of the acid is to weaken the O-I bond.

Painter and Soper (41) have studied the iodination of phenols by iodine in solutions buffered by weak acids. They found that the speed of iodination varied directly as the concentration of phenol and inversely as the square of the iodide concentration. They also found that the speed varies directly as the concentration of the buffer acid. They therefore concluded that iodination is by HOI on phenol and AcOI on phenoxide ion or by I^+ on phenoxide ion and AcOI on phenoxide ion. These results should be compared with those of Wilson and Soper (36), who found that in acid solutions weak acids did not catalyze bromination of o-nitroanisole and benzene by bromine water. This indicates that bromine is not hydrolyzed to an appreciable extent and inappreciable amounts of acylhypobromites are formed in such acid solutions. This difference was also predicted by Soper

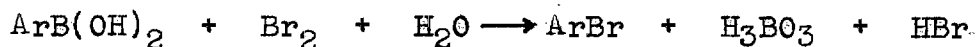
41. B. S. Painter and F. G. Soper, J. Chem. Soc., 1947, 342.

and Smith (30) on the basis of the ionization and hydrolysis constants of the halogens.

Survey of electrophilic displacement of groups other than hydrogen.--A few of the electrophilic displacements of

groups other than hydrogen have been studied rather extensively. Some of these will be reviewed in this section following which a number of typical reactions will be listed.

Kuivila and Esterbrook (42) have studied the following reaction:



They have found that the reaction in aqueous acetic acid is first order with respect to bromine (the stoichiometric bromine concentration must be corrected for that which is converted to tribromide ion by bromide ion) and first order with respect to the benzene boronic acid. The reaction is accelerated by salts of strong acids and salts of weak acids exert pronounced catalytic effects. The rate of the reaction can be satisfactorily expressed by:

$$dx/dt = [k_o + k_c(\text{base})][\text{ArB(OH)}_2][\text{Br}_2]_{\text{actual}}.$$

Kuivila and Hendrickson (43) have studied the same reaction with substituted boronic acids. From this study they have concluded that cleavage of the C-B bond is not important in

42. H. G. Kuivila and E. K. Esterbrook, J. Am. Chem. Soc., 73, 4629 (1951).

43. H. G. Kuivila and A. R. Hendrickson, J. Am. Chem. Soc., 74, 5068 (1952).

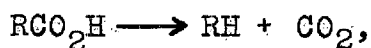
the rate determining step and that the critical step is the formation of the pictured intermediate.



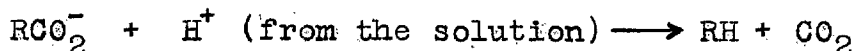
Schubert and Gardner (44) have studied the decarboxylation of 2,4,6-trihydroxybenzoic acid over a wide range of acid concentration¹ (0 to 58.5 per cent perchloric acid). The reaction was followed spectrophotometrically (u. v.) which had the advantage of enabling the workers to determine the actual concentrations of RCOOH and RCOO^- in the reaction medium. In water at 10^{-4} molar, the acid was found to be practically completely ionized to its anion and no decarboxylation was observed, hence the rate of the unimolecular decomposition of RCO_2^- is negligible if not zero. They found no evidence of an acid catalysis of the type suggested for the decarboxylation of the alkyl benzoic acids, i.e., decarboxylation by way of the conjugate acid. The data which they found are consistent with a mechanism of either an $\text{S}_{\text{N}}1$ type unimolecular rearrangement of undissociated RCO_2H ,

44. W. M. Schubert and J. D. Gardner, J. Am. Chem. Soc., **75**, 1401 (1953).

¹This reaction has also been studied by H. Schenkel and M. Schenkel-Rudin, Helvetica Chimica Acta, **31**, 514 (1948), and by B. R. Brown, W. W. Elliot and D. L. Hammick, J. Chem. Soc., **1951**, 1384.



or a bimolecular transfer of a proton from the solution to the carboxylic acid anion to give the products in either a concerted displacement or a multiple step reaction.



Schubert, Donohue, and Gardner (45, 46) have studied the kinetics of the decarboxylation of trialkylbenzoic acids in 82-100 per cent sulfuric acid. The rate was measured by the evolution of carbon dioxide. The rate was found to go through a maximum at 85 per cent sulfuric acid. They have also measured spectrophotometrically the concentrations of and the equilibria between the un-ionized aromatic acid and its conjugate acid and between the conjugate acid and the acylium ion. Their data indicate that the reaction does not proceed by the Hammett unimolecular mechanism, i.e., a unimolecular decomposition of the conjugate acid into the products. They believe that catalysis by molecular sulfuric acid is the predominant process in greater than 80 per cent sulfuric acid, but that general acid catalysis cannot be ruled out, since

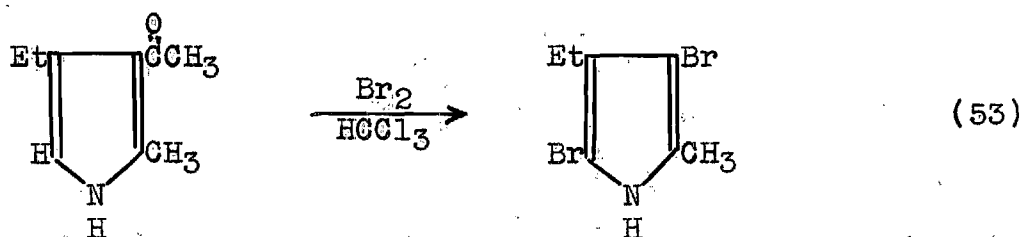
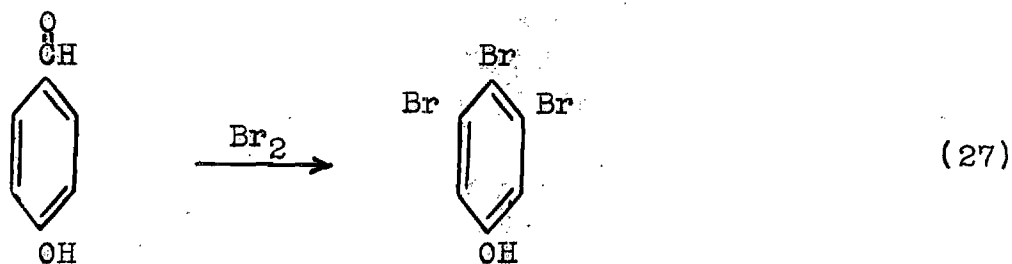
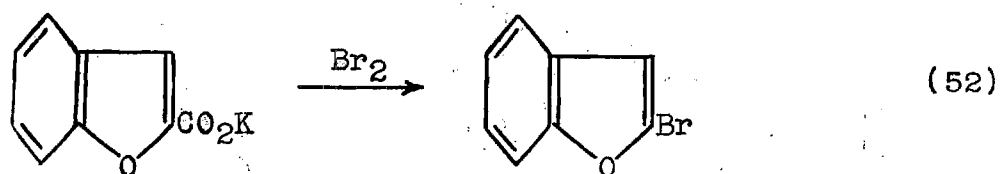
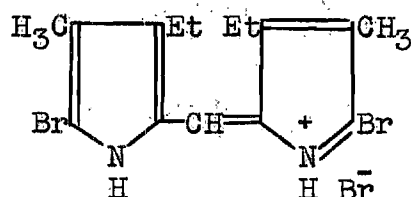
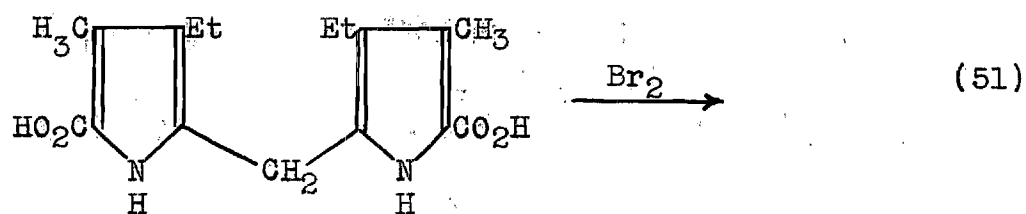
45. W. M. Schubert, J. Am. Chem. Soc., 71, 2639 (1949).

46. W. M. Schubert, J. D. Gardner and J. Donohue, J. Am. Chem. Soc., 76, 9 (1954).

below 80 per cent sulfuric acid, the participation of oxonium ion catalysis appears to be discernible. From the work of Stevens, Pepper, and Lounsbury (47), it seems that the rate determining step is in the breaking of the C-C bond. This is demanded because of the 3-4 per cent isotope effect (between C^{12} and C^{13}) that they observed in the decarboxylation of mesitoic acid.

Stevens, Pepper, and Lounsbury (47) have studied the decarboxylation of anthranilic acid and found that the aqueous decomposition can be acid catalyzed but that when the concentration of mineral acid approaches that of the anthranilic acid, the rate is decreased. Anthranilic acid with C^{13} in the carboxyl group was found to decompose at the same rate as that with C^{12} in the carboxyl group by a mass spectrograph analysis of the carbon dioxide that was evolved. They interpret the mechanism of this decarboxylation as being a bimolecular electrophilic substitution, with the attack of a proton, probably on the α -carbon of the zwitter ion, for the rate controlling step.

47. W. H. Stevens, J. M. Pepper and M. Lounsbury, Canadian Journal of Chemistry, 30, 529 (1952).

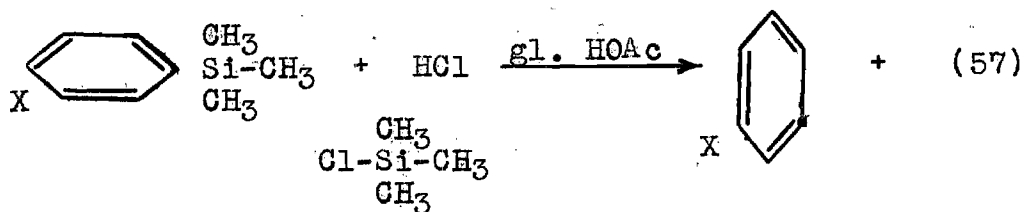
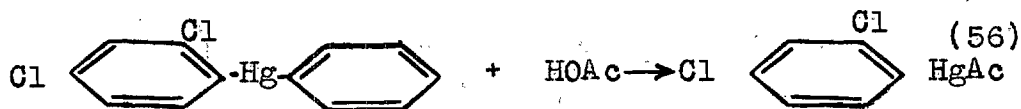
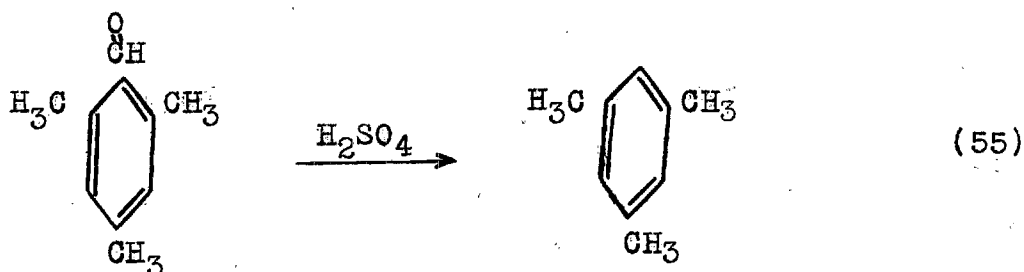
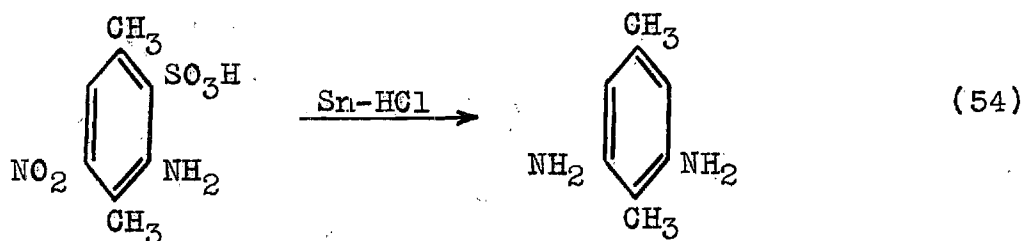
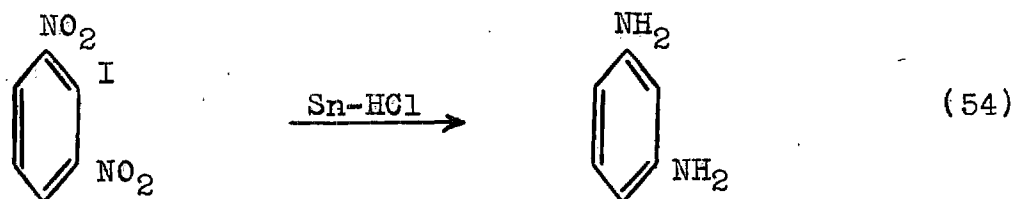


51. H. Fischer, P. Halbig and B. Walach, Ann., 452, 268 (1927).

52. R. Stoermer and G. Calov, Ber., 34, 770 (1901).

53. H. Fischer and R. Bauml, Ann., 468, 58 (1929).

Electrophilic displacements by hydrogen:



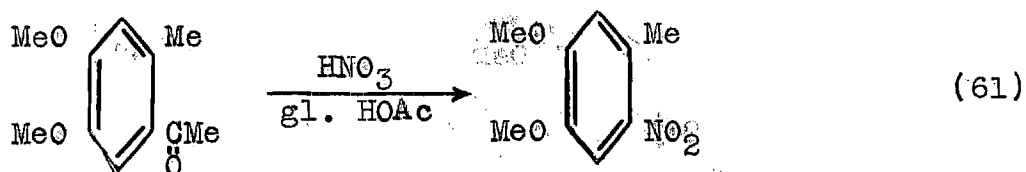
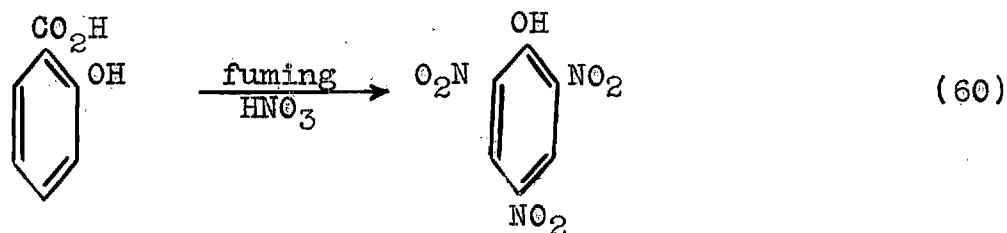
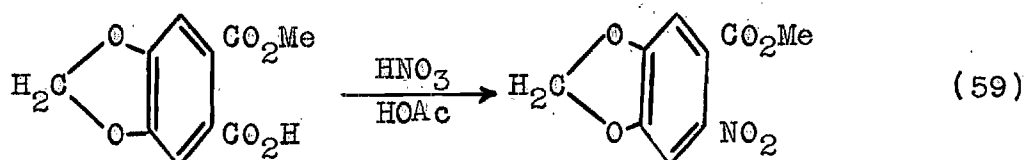
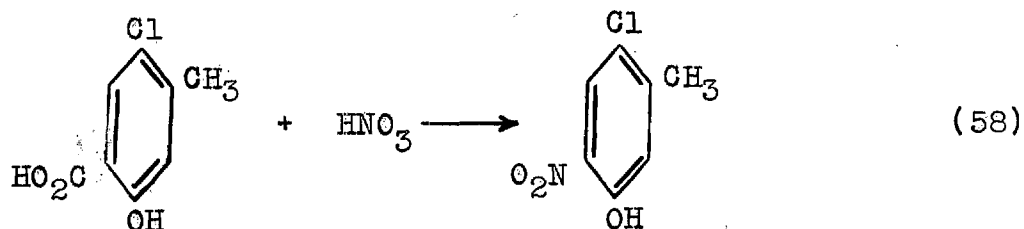
54. M. J. J. Blanksma, Rec. trav. chim., **24**, 320 (1905).

55. W. M. Schubert and R. E. Zahler, J. Am. Chem. Soc., **76**, 1 (1954).

56. A. H. Corwin and M. A. Naylor, J. Am. Chem. Soc., **69**, 1004 (1947).

57. R. A. Benkeser and A. Torkelson, J. Am. Chem. Soc., **76**, 1252 (1954).

Electrophilic displacements by a nitro group:

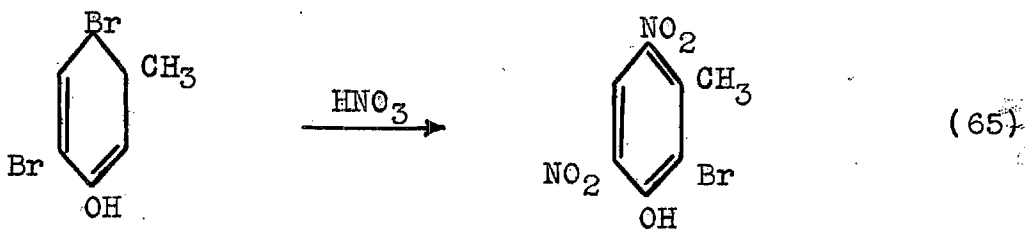
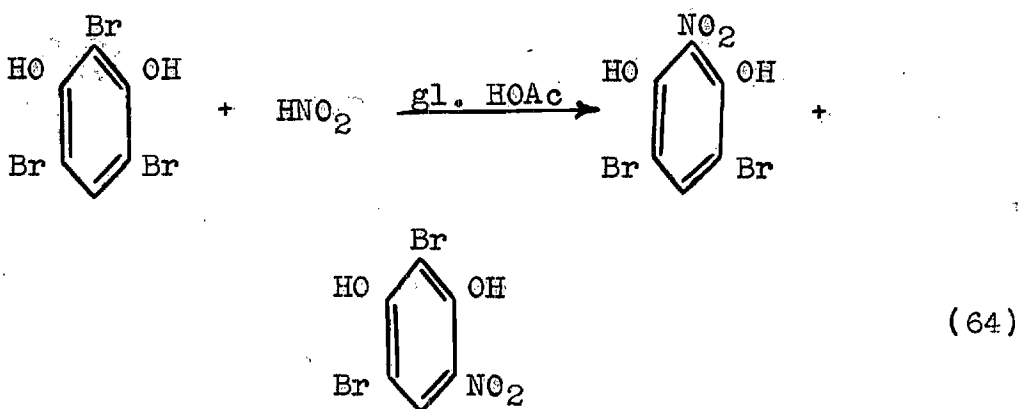
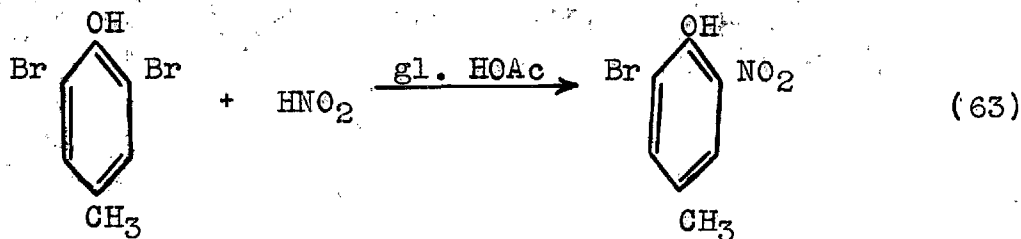
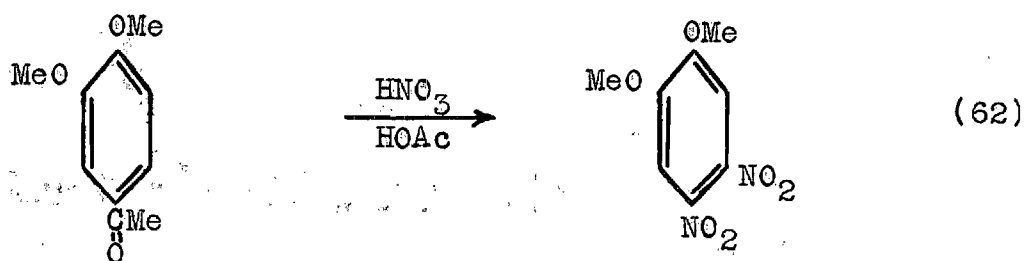


58. I. R. Gibbs and P. W. Robertson, J. Chem. Soc., 105, 1885 (1914).

59. V. J. Harding, J. Chem. Soc., 99, 1585 (1911).

60. D. Nightingale, Chem. Rev., 40, 117 (1947).

61. V. J. Harding and C. Weizmann, J. Chem. Soc., 97, 1126 (1910).

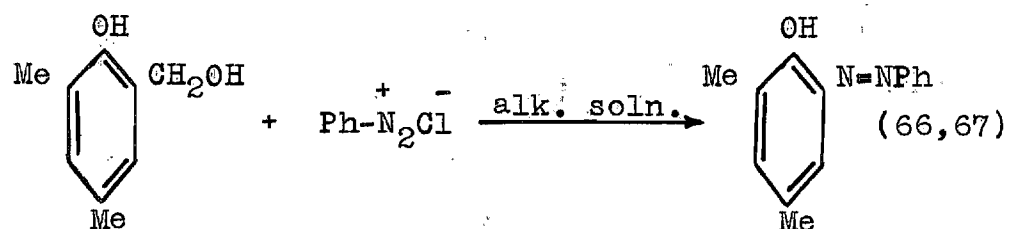
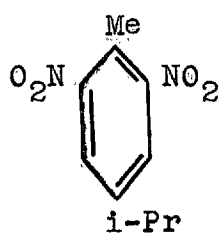
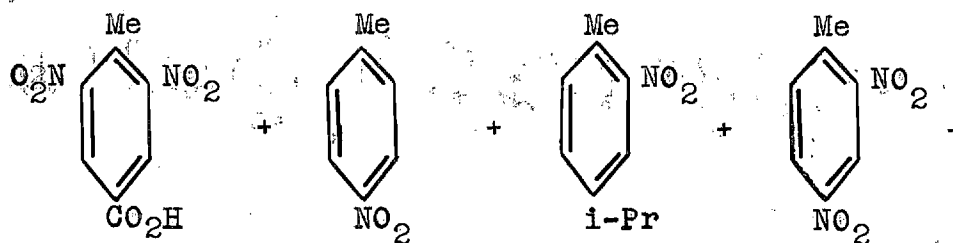
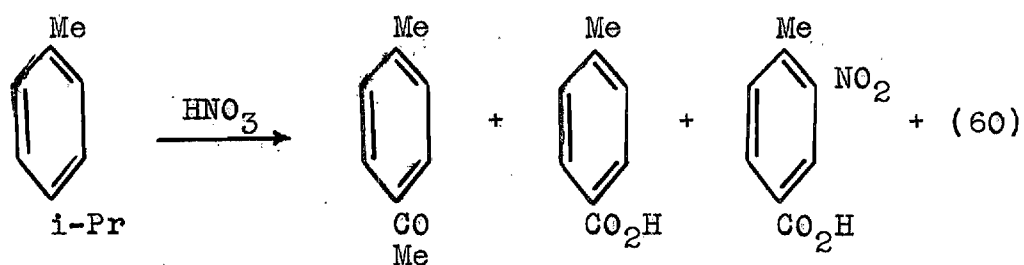


62. H. Thomas and W. Siebeling, Ber., 44, 2115 (1911).

63. T. Zinke, J. prakt. chem., (2), 61, 561 (1900).

64. G. Dahmer, Ann., 333, 346 (1904).

65. P. W. Robertson and H. V. Briscoe, J. Chem. Soc., 101, 1964 (1912).



66. E. Marder and I. W. Ruderman, J. Am. Chem. Soc., **73**, 5475 (1951).

67. E. Ziegler and G. Zigeuner, Monatsh, **79**, 42 (1948).

CHAPTER III

SOURCE, PREPARATION, AND PURIFICATION OF MATERIALS

Acetic acid.--The acetic acid used for all preparations and for kinetic Runs 6-10 was DuPont's C. P. glacial acetic acid used without further purification.

The acetic acid used in the solvents for Runs 11-17, acetic acid-(a), was prepared by first distilling the glacial acetic acid and collecting the fraction boiling between 115-116.8°. The freezing point of this mixture was observed to be 15.2° (corr.) indicating that the solution was ca. 0.368 molal in water. Sixty five milliliters of 97 per cent acetic anhydride (ca. 0.68 moles) was added to 3500 ml. (3670 g.) of the distillate and the mixture was refluxed for four hours. The corrected freezing point of this mixture was observed to be 16.2° indicating a solution ca. 0.105 molal in water.

The acetic acid for all other kinetic runs, acetic acid-(b), was purified by the method of Bradfield and Orton (68). Four liters of the DuPont acetic acid was refluxed for two hours over 80 g. of chromium trioxide. The solution was then distilled through a three foot, 1 3/8 inch bore, vacuum jacketed column packed with 3/16 inch glass helices,

68. A. E. Bradfield and K. J. Orton, J. Chem. Soc., 1924, 960.

and the fraction boiling between 116-117° was collected (ca. 3 liters). The solution was then analyzed for water content by titration with Karl Fischer Reagent (69). The Karl Fischer Reagent was run from a bottom filling buret having an extended tip and equipped with a reservoir so arranged that only air which had been passed through a Drierite and silica gel trap could be admitted into the system. The 20 ml. samples of acetic acid were titrated in 100 ml. volumetric flasks to an end point intermediate between the dark brown iodine color and the canary yellow of the aqueous solution. It was observed that acetic acid gave a false end point upon the addition of less than 1 ml. of reagent. This end point disappeared upon the addition of more reagent and the true end point then appeared. All acetic acid prepared by this method contained less than 1 per cent water and the amount of water was known to within 0.01 per cent (the variation in concentration that would result from an uncertainty of 0.5 ml. of Karl Fischer Reagent).

Acetic anhydride.--Acetic anhydride of unknown purity and source was used without further purification.

Acetonitrile.--Eastman Kodak, White Label, acetonitrile was used without further purification.

69. J. Mitchell, Jr. and D. M. Smith, Aquametry, New York, Interscience Publishers, 1948.

Barium hydroxide.--Baker's Analyzed C.P. barium hydroxide octahydrate was used to prepare all solutions. To prepare the standard base, 253 grams of the barium hydroxide was dissolved in 4 liters of hot distilled water. The barium carbonate formed in the process was filtered and the clear solution poured into a polythene bottle which had been previously flushed with nitrogen and was equipped with a soda lime trap. The base was standardized against standard hydrochloric acid.

Bromine.--Merck, Reagent Grade, bromine was used for all preparations and for Runs 1-12 without further purification.

For kinetic Runs 13-44, Baker's Analyzed C.P. bromine was used without further purification. For all other kinetic runs, this bromine was purified by the following procedure¹: Six hundred milliliters of bromine was mixed with 60 g. of potassium bromide and refluxed for three hours. The mixture was then distilled, the first 100 ml. of distillate was rejected, and the next 350 ml. of bromine, boiling between 57-58°, was collected and used without further purification.

Chloroacetic acid.--Matheson chloroacetic acid was used without further purification.

¹c.f. Kuivila and Esterbrook (42).

Chloroform.--Commercial chloroform was used without further purification.

Chromium trioxide.--Merck's chromic acid (chromium trioxide) was used in the purification of acetic acid.

3,5-Dibromo-4-hydroxybenzoic acid (DBPHBA)

Preparation.--Twenty five grams (181 m. moles) of p-hydroxybenzoic acid was suspended in 175 ml. of glacial acetic acid, heated on a steam bath, and 50 ml. (0.94 moles) of bromine was added dropwise. The product commenced precipitating when about half of the bromine had been added. Heating was continued for two hours, after which the mixture was cooled and filtered. The crude DBPHBA was transferred to a 500 ml. flask and dissolved in 290 ml. of refluxing glacial acetic acid and the solution filtered through a heated sintered glass funnel. Upon cooling to room temperature, the DBPHBA crystallized out in fine white crystals. These were filtered and washed with two 20 ml. portions of glacial acetic acid, sucked almost dry, and finally dried in a vacuum oven for six hours at 50°. The yield was 29.3 grams (55 per cent) of a product melting at 275-276° ¹. A neutralization equivalent of about 150 (theoretical value 148) at a pH of 11 was obtained in a dioxane-water solvent,

¹Unless otherwise stated, all melting points were obtained from an Anschütz thermometer in an electrically heated Thiele type melting point apparatus.

indicating that both of the acid hydrogens were neutralized at the same time. The accuracy of the determination was limited by instrumentation difficulties.

Purification Method A.--Dibromo-p-hydroxybenzoic acid purified by this method will be designated as DBPHBA_a.

Ten grams of DBPHBA (ca. 34 m. moles) were dissolved in ca. 350 ml. of ether. To this mixture 20 ml. of aniline (ca. 200 m. moles) and 350 ml. of normal pentane were added. Upon shaking the mixture, a voluminous precipitate formed. The mixture was cooled in ice water, then filtered, and the precipitate washed with four 25 ml. portions of a 1:4 ether-pentane mixture. The crude anilide thus formed was treated with ca. 55 ml. of 6 N sulfuric acid followed by 25 ml. of water. The white paste thus obtained was filtered with difficulty and then washed with 100 ml. of ice water. The precipitate (still strongly acid) was then dissolved in 100 ml. of ether. The water phase which separated was removed. The ether phase was washed with three 10 ml. portions of water to remove any sulfuric acid. The ether was allowed to evaporate and the DBPHBA recovered (8.8 g.). This DBPHBA was then recrystallized once from 65 ml. of glacial acetic acid and then from 65 ml. of acetonitrile. The crystals of DBPHBA_a were dried in a vacuum oven at 50° for eight hours. The final yield was 4 g. of a product melting at 275°. The melting point

indicated no improvement of product. A sample was sent away for analysis¹ with the following results:

	Found	Calculated (for $C_7H_4O_3Br_2$)
Carbon	29.18%	28.4 %
Hydrogen	1.82	1.35
Bromine	52.98	54.1

Loss of weight on drying - 1.53%

Purification Method B.--Dibromo-p-hydroxybenzoic acid purified by this method will be designated as DBPHBA_p. Eleven grams of DBPHBA was dissolved in 500 ml. of 1 N sodium hydroxide solution and refluxed for four hours to saponify any acetate esters present. The solution was then acidified with 70 ml. of 85 per cent phosphoric acid and the solution filtered while hot. The precipitate was washed several times with water until it no longer smelled of phenol. The precipitated DBPHBA was then shaken successively with 80 ml. and 40 ml. portions of n-pentane to remove any residual phenol. The DBPHBA was then suspended in 100 ml. of water and neutralized to a phenolphthalein end point with about 250 ml. of a ca. 0.1 N solution of barium hydroxide. The barium hydroxide thus formed by the neutralization of both acid hydrogen ions of DBPHBA was precipitated by the addition of 1000 ml. of 95 per cent ethanol. The salt was filtered and washed with 50 ml. of ethanol. The dry barium salt was dissolved in 400 ml. of

¹Clark Microanalytical Laboratory, Urbana, Illinois.

warm water and DBPHBA was precipitated by the addition of 5 ml. of 12 N hydrochloric acid. The precipitate was filtered and washed and finally dried in a vacuum oven for four hours at 50°. The yield of DBPHBA was 6.7 g. The DBPHBA was recrystallized from 200 ml. of acetonitrile. The yield was 4.6 g. As the final step in the purification, 2.0 g. of DBPHBA was recrystallized from 1 liter of chloroform. It was necessary to cool the chloroform solution on dry ice to obtain a satisfactory yield. The yield of DBPHBA_p was 1.46 g. of a product melting at 275-276°. Since the very high melting point of the compound make the melting point a poor index of purity, a sample was analyzed¹ with the following results:

	Found	Calculated (for C ₇ H ₄ O ₃ Br ₂)
Bromine	54.34%	54.1%

Loss of weight on drying - none

This is apparently an improvement of the product purified by Method A.

3,5-Dibromo-2-hydroxybenzoic Acid (DBSA)

Preparation.--3,5-Dibromosalicylic acid was prepared by the method of Earle and Jackson (12). In a typical preparation, 25.0 g. (181 m. moles) of salicylic acid was suspended in 125 ml. of glacial acetic acid in a 250 ml. round bottom flask equipped with a reflux condenser which was in turn

¹Clark Microanalytical Laboratory, Urbana, Illinois.

attached to a vapor absorption spray. Fifty milliliters of bromine (0.94 moles) were added and the mixture heated on a steam bath for three hours. The mixture was allowed to cool and the resulting crystals were removed by vacuum filtration in a sintered glass funnel. The crystals were washed with two 20 ml. portions of glacial acetic acid and then recrystallized from 100 ml. of boiling glacial acetic acid. The yield was 39.97 g. (74.6 per cent yield) of DBSA melting at 226-227°. The product was then dissolved in 75 ml. of boiling glacial acetic acid and filtered through a heated sintered glass funnel. Cooling of the filtrate to room temperature caused crystallization of DBSA. The crystals were filtered, washed once with 10 ml. of glacial acetic acid, sucked dry, and finally dried for six hours in a vacuum oven at 50°. A pH titration of DBSA was not too successful because of instrumentation difficulties, however only one neutralization equivalent, in the vicinity of 290 (theoretical value 296), could be obtained.

Purification Method A.--Some of the DBSA was further purified by the following process and is designated by DBSA_a. Thirty grams of DBSA were refluxed for four hours in 1500 ml. of 1 N potassium hydroxide to saponify any phenol acetates that might have been formed in the bromination step. The mixture was neutralized with 200 ml. of 85 per cent phosphoric acid and filtered while hot. The

precipitate of DBSA was washed several times with water, sucked dry, and then shaken successively with 100 ml. and 50 ml. portions of n-pentane to extract any phenol present. The dried DBSA was suspended in 300 ml. of water and titrated to a phenolphthalein end point with barium hydroxide solution. A new solid was formed in this process which was indicated by the titer of base to be the barium salt formed by the neutralization of one acid hydrogen per molecule of DBSA. The barium salt suspension was filtered and washed with water, then triturated with a mixture of 20 ml. of 12 N hydrochloric acid and 100 ml. of water. The DBSA thus regenerated was filtered, washed three times with water and allowed to dry in air. The next step in the purification was a recrystallization from 100 ml. of acetonitrile. To aid in the crystallization of the product, the solution in acetonitrile was chilled on dry ice, whereupon 18 g. of DBSA was recovered. The DBSA was finally dried under vacuum in an Abderhalden apparatus containing phosphorous pentoxide and heated by refluxing acetone. The melting point of the final product was 228-230°. It is believed that the procedure involving refluxing with potassium hydroxide and formation of the barium salt could probably have been omitted without altering the purity of the product.

2,4-Dibromophenol (2,4-DBP).--Ten grams of 2,4-dibromophenol (Eastman Kodak, White Label, m.p. 32.5-36°) was

dissolved in 10 ml. of chloroform and recrystallized at -40° . The yield was 6.1 grams of product melting at $35.5-36.5^{\circ}$. One more crystallization of the 2,4-DBP from chloroform at -40° gave a product melting at $36.2-37^{\circ}$ (recorded m.p. 40° (70)).

2,6-Dibromophenol (2,6-DBP).--Thirteen grams of 2,6-dibromophenol (Eastman Kodak, White Label, m.p. $54-56^{\circ}$) was distilled at 18 mm. pressure. The distillate was dissolved in 50 ml. of chloroform and cooled in a dry ice-acetone bath until 2,6-DBP crystallized out. Four grams of 2,6-DBP was obtained after filtration and drying under vacuum at room temperature. The melting point of the product was $56-57^{\circ}$ (recorded m.p. $56-57^{\circ}$ (70)).

Dioxane.--Dioxane from an unknown commercial source was purified according to the method described in Fieser's Lab Manual (71).

Hydrogen bromide.--Matheson's C.P. hydrogen bromide was passed into distilled water to form a solution of ca. 1.5 N. The exact concentration of the solution was determined by titration with standard barium hydroxide solution.

70. I. M. Heilbron and H. M. Bunbury, Dictionary of Organic Compounds, New York, Oxford University Press, 1943.

71. L. F. Fieser, Experiments in Organic Chemistry, Part II, 2nd Edition, Boston, Mass., D. C. Heath and Co., 1941, p368.

This titration has been checked by titrating the bromide ion with standard silver nitrate solution using eosin indicator.

p-Hydroxybenzoic acid.--Matheson and Eastman Kodak, White Label p-hydroxybenzoic acid (m.p. $213-214^{\circ}$) was recrystallized from boiling water to give a product melting at $213.5-214.5^{\circ}$ (recorded m.p. $213-214^{\circ}$ (70)).

Karl Fischer Reagent.--Sargent's Karl Fischer Reagent was used for all water in acetic acid determinations. Eimer and Amend's Karl Fischer Reagent was not satisfactory because it contained a red dye that obscured the visual end point.

Lithium bromide.--Lithium bromide solutions were prepared from Coleman and Bell's C.P. lithium bromide and distilled water. The exact concentration of the solutions was obtained by titration of the bromide ion with standard silver nitrate solution and eosin indicator as per procedure in Kolthoff and Sandell (72a).

Methanol.--Anhydrous methanol was prepared by allowing 10 g. of magnesium to react with 1500 ml. of methanol, refluxing for two hours, then distilling the mixture and

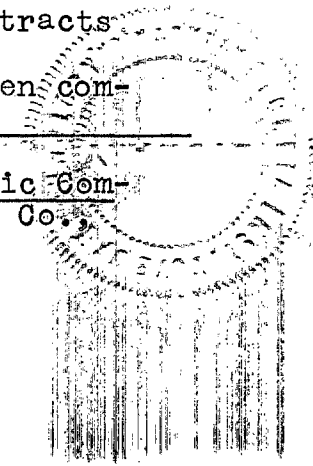
72a. I. M. Kolthoff and E. B. Sandell, Textbook of Quantitative Inorganic Analysis, (3rd Edition), New York, The Macmillan Company, 1952, p. 544.

collecting the distillate boiling at 63°.

A standard water in methanol solution was prepared by weighing 3.9555 g. of distilled water at 20.0° into enough anhydrous methanol to make 250 ml. of solution at 20.0°.

Methyl 2,4-dibromophenyl ether.--The ether was prepared according to the procedure recorded in Hickinbottom (73) for the preparation of phenetole. Fifty grams of freshly distilled 2,4-dibromophenol was mixed with a solution of sodium methoxide which had been prepared by dissolving 4.3 g. of sodium in 75 ml. of methanol. Ten milliliters of methyl iodide was then added and the solution refluxed. After about twenty minutes, 10 ml. more of methyl iodide was added and the solution refluxed until neutral to litmus (about four hours). The solvent methanol was distilled from the solution and the residue dissolved in 50 ml. of ether and 25 ml. of water. The water phase, containing the sodium iodide was separated and discarded. The ether phase was washed with two 20 ml. portions of a 10 per cent sodium carbonate solution and with a 20 ml. portion of water. The combined water and sodium carbonate extracts were then washed with 10 ml. of ether which was then com-

73. W. J. Hickinbottom, Reactions of Organic Compounds, (2nd. Edition), London, Longmans Green and Co., 1948, p. 91.



bined with the main ether phase. The combined ether phases were dried over Drierite, decanted into an evaporating dish, and the ether allowed to evaporate. The yield was 51.3 g. (97 per cent). The crude ether was recrystallized from 60 ml. of methanol, yielding 44.7 g. of material melting at 62-63° (corr.), (recorded value 61.3° (14)). This material was then further purified by recrystallization from 120 ml. of n-hexane. The yield from this step was 35.4 g. (69 per cent overall yield) of material melting at 62.6-63.6° (corr.).

Methyl iodide.--Eastman Kodak, White Label, methyl iodide was used without further purification.

Perchloric acid.--Merck's Reagent 70 per cent perchloric acid was used without further purification. The exact concentration at various times was determined by titration against a standard base.

Phenol.--Merck's U.S.P. phenol was used without further purification.

Phosphoric acid.--Baker's N.F. 85 per cent phosphoric acid was used.

Potassium bromate.--Merck's Reagent potassium bromate was used without further purification.

Potassium bromide.--Merck's Reagent potassium bromide was used without further purification.

Potassium dichromate.--Fifty grams of Baker and Adamson's Reagent potassium dichromate was recrystallized from 100 ml. of boiling water. The crystals obtained were then recrystallized again from 50 ml. of boiling water and dried under vacuum in an Abderhalden apparatus containing phosphorous pentoxide and heated by refluxing bromobenzene.

This procedure was used after Run 43 when it was discovered that the dichromate that had been used previously was sodium and not potassium. An error was initially suspected when some sodium thiosulfate solution that had been standardized against the dichromate was used by a colleague, Arthur M. Dowell, in a titration against some standard sodium thiocresylate. The error was confirmed when a standard sodium bromate solution was prepared and used to standardize the thiosulfate. The concentration thus indicated for the thiosulfate was almost 10 per cent greater than that previously used. This would be the result if the thiosulfate had been standardized with sodium dichromate and calculations made on the basis of potassium dichromate. The error was finally confirmed by a flame test on the suspect dichromate and sample of known potassium dichromate. The former gave a very positive sodium test. The original source of

this sodium dichromate is not known. The concentrations of all sodium thiosulfate solutions were then corrected and the results of all kinetic runs recalculated. A marked improvement in the kinetic data was then noted.

Potassium hydroxide.--Baker's Analyzed Reagent potassium hydroxide was used.

Potassium iodide.--Merck's Reagent potassium iodide was used without further purification.

Potassium perchlorate.--Four grams of potassium hydroxide was dissolved in 30 ml. of distilled water and neutralized with 8 ml. of 11.7 M perchloric acid. The white precipitate was removed by filtration, washed with distilled water until neutral to litmus, and dried over phosphorous pentoxide.

Sodium Thiosulfate.--A stock 0.1 N solution was prepared as per Kolthoff and Sandell (72b), using Merck's Reagent sodium thiosulfate decahydrate. The solution was diluted to ca. 0.02 N before using and was then standardized with a standard potassium dichromate solution. In a typical standardization, 10.00 ml. of 0.0694 N potassium dichromate solution, 10 ml. of water, 2 ml. of concentrated C.P.

72b. I. M. Kolthoff and E. B. Sandell, op. cit., p. 592.

hydrochloric acid, and 4 ml. of 1 N potassium iodide were titrated with the thiosulfate solution to the complete disappearance of the yellow tint due to iodine. No starch indicator was found necessary to obtain reproducible results.

Salicylic acid.--Matheson's salicylic acid (m.p. 157-159°) was recrystallized from boiling water to give a product melting at 157.5-159° (recorded m.p. 159° (70)). Salicylic acid will slowly decarboxylate in boiling water, consequently, this method of purification must be used cautiously.

2,4,6-Tribromophenol (TBP).--Tribromophenol was prepared according to the method of Adkins, McElvain, and Klein (74). Ten grams of phenol was dissolved in 500 ml. of water and bromine vapor was drawn into the solution by means of a water aspirator until precipitation ceased and the solution became slightly yellow (ca. 17 ml. of bromine was required). The precipitate was filtered and washed with an acidified solution of sodium sulfite to remove any unreacted bromine, then washed with water. The precipitate was dissolved in 750 ml. of alcohol at reflux and tribromophenol was precipitated from the solution by the addition of 1500 ml. of water. The TBP was filtered and dried yielding 20.35 g. of product melting at 92-93° (recorded m.p. 94° (70)).

74. H. Adkins, S. M. McElvain, and M. W. Klein, Practice of Organic Chemistry, 3rd Edition, New York, McGraw-Hill Book Company, Inc., 1940, p. 81.

CHAPTER IV

EXPERIMENTAL TECHNIQUES FOR KINETIC MEASUREMENTS

Bromine Consumption Method

General.--It was initially assumed that the reaction under investigation was first order in bromine and first order in hydroxybenzoic acid and integrated second order rate constants were calculated on this basis. The rate equation used was the conventional second order equation

$$dx/dt = k_{app}(a-x)(b-x)$$

in which a is the initial molar concentration of substrate, b is the initial molar concentration of bromine (stoichiometric concentration), and x is molar concentration of substance reacted (or product formed) at time t . The integrated form of this equation that was used to calculate the integrated rate constants is

$$k_{app} = \frac{2.303}{(b-a)t} \log \frac{a(b-x)}{b(a-x)}.$$

The units of k_{app} are (liter)(mole⁻¹)(minute⁻¹) or (l./mol.min.). All of the rate constants cited in the tables were calculated on this basis.

Throughout the course of the experimental work, two general techniques for measuring the kinetics of the reactions were used. In Runs 1-10 and Run 46, the reactions were carried out in a single flask and aliquots were withdrawn at noted times and the unconsumed bromine titrated after stopping the reaction. In the remainder of the runs, the reaction was run concurrently in a number of flasks. The reaction was then stopped in individual flasks at noted times and the contents analyzed for unconsumed bromine.

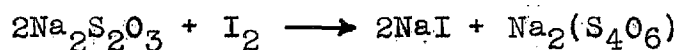
In all cases, analysis for bromine was by reaction with potassium iodide followed by titration with standard sodium thiosulfate solution of approximately 0.02 N. Initially, chloroform was used as an indicator in this reaction but it was soon found that no indicator was necessary, and that the disappearance of the yellow iodine color could be observed reproducibly in all of the solvents used. In most cases, a 50 ml. buret was used to deliver the sodium thiosulfate solution; however, in cases where the titer was less than ten milliliters, a 10 ml. buret was generally used.

For those reactions carried out at temperatures less than 10°C., the solutions were thermostated in an "Aminco" low temperature bath capable of maintaining a temperature to within 0.5°C. All other runs were thermostated in a Sargent Constant Temperature Water Bath, S-84805, equipped with a Sargent Mercurial Thermoregulator having a sensi-

tivity of $\pm 0.01^{\circ}\text{C}$. The thermoregulator was satisfactory in the winter months when the humidity was comparatively low, however in warm weather and periods of high humidity, it had a very annoying tendency to collect condensate and cease functioning. For Runs 64 and following, the Sargent thermoregulator was replaced by an H. B. Instrument Company hermetically sealed, hydrogen filled mercurial thermoregulator having a sensitivity of $\pm 0.001^{\circ}\text{C}$. The temperature was observed on a $0-100^{\circ}\text{C}$. thermometer calibrated in tenths of a degree. The thermometer was checked against a Bureau of Standards thermometer at 20.0°C . and no discrepancy could be noted. The bath temperature was observed to stay constant within 0.02° .

Sample withdrawal method.--Runs 1-4 in dioxane-water solution, Run 5 in ethanol-water solution, and Runs 6-10 in acetic acid-water solution were all run by this technique. In a typical run of this type, 80 ml. of a solution of DBSA in the desired solvent (mixed by volume per cent unless otherwise stated) was placed in a low actinic 100 ml. volumetric flask which was then thermostated at the desired temperature. Twenty milliliters of a bromine solution in the same solvent was then added, the time noted, and the flask shaken vigorously. The initial bromine concentration (as well as any possible side reaction with the solvent) was determined by adding 20.0 ml. of the bromine solution

to 80.0 ml. of the same solvent, and titrating the bromine in aliquots withdrawn from the solution. To follow the reaction, aliquots of 10.0 or 15.0 ml. were withdrawn at noted times and placed in flask containing 2 ml. of 1 N potassium iodide and 10 ml. of distilled water. The mixture was then titrated with standard sodium thiosulfate solution.



Multiple flask method.--In general, in this method a stock solution of the desired solvent (calculated as weight per cent unless otherwise stated) was prepared and with this solvent 250 ml. of a stock solution of the compound being brominated was prepared. In Runs 11-50, 20.0 ml. of the solution was placed in each of a number of 50 ml. volumetric flasks, and in the remaining runs 40.0 ml. of the solution was placed in each of a number of 100 ml. volumetric flasks. A like volume of the solvent was also placed in each of several volumetric flasks to serve as blanks. In order to avoid any possible photo-catalysis of the reaction, the solutions were always placed in either black painted flasks or red low actinic glassware.

Two different methods of preparing a bromine solution were used. The first method, used on Runs 11-17, consisted of first preparing a saturated bromine water solution and

mixing this with distilled water and glacial acetic acid to prepare a bromine solution of the same solvent concentration as that being used in the reaction. The second, and more accurate method, consisted of pipetting a small volume of liquid bromine into some of the stock solvent solution previously prepared. It was found, for example, that if 0.2 ml. of bromine were added to 100 ml. of solvent, the resulting solution was ca. 0.04 M in bromine. The bromine concentration in the reaction could thus be regulated with reasonable accuracy. When the bromine solution had been prepared in either case, then 5.0 or 10.0 ml. of it was added at noted times to each flask that had been previously prepared so as to make respectively 25 or 50 ml. of reaction mixture.

The method of stopping the reaction varied somewhat as experience showed to be necessary. On Runs 11-36, at the proper time the reaction mixture was poured into an Erlenmeyer flask containing about 10 ml. of water and 4 ml. of 1N potassium iodide solution. The reaction flask was then rinsed with a 0.05 N KI solution. On the remaining runs, the potassium iodide was injected into the reaction flask with a syringe because it was felt that this action could be timed with more precision and the force of the injected stream would cause more rapid mixing and quicker stopping of the reaction. It was feared that a slow mixing

of the potassium iodide solution might momentarily speed up the bromination reaction by increasing the water concentration since this has been shown to greatly accelerate the speed of the bromination.

It was observed very early that the acidified acetic acid solvents used caused a very rapid oxidation of iodide ion to iodine and as a result it was almost impossible to obtain an end point in the titration. Consequently, it was decided to neutralize most of the acid in these solutions with potassium hydroxide solution as soon as possible after the potassium iodide had been added. In Runs 12-38 this was done by pouring 10 ml. of 4 N potassium hydroxide solution into the titration flask as soon as possible after the reaction had been stopped. There was a considerable time lag in this technique; therefore, in order to minimize the oxidation of iodide ion to iodine, on Runs 39 and all following, 10 ml. of 4 N potassium hydroxide was injected into the reaction flask by means of a syringe immediately following the injection of the potassium iodide solution. The mixture was then poured into 10 ml. of distilled water in an Erlenmeyer flask, the reaction flask rinsed with water, and the solution titrated with sodium thiosulfate solution.

It seemed possible that the iodine (or triiodide ion) formed in stopping the reaction might itself continue

to react with the remaining hydroxybenzoic acid in the solution. In order to test this hypothesis, Run 38 was made in which the stopped reaction mixtures were allowed to stand for varying periods of time before titration. As shown in Table 1, this had no significant effect on the titer of thiosulfate required, indicating that iodine (or tri-iodide ion) did not react under the conditions of the experiment.

Carbon dioxide evolution method.--It was found experimentally that if 80 per cent acetic acid solvent was saturated with carbon dioxide, any carbon dioxide liberated by a reaction taking place in this solvent could be discharged from the solution if the mixture was shaken vigorously enough. A system was devised, as follows, to measure the carbon dioxide evolved from DBPHBA by the action of bromine:

The reaction was carried out in a 500 ml. "angle type", standard taper three neck flask. In order to flush the system with carbon dioxide, one of the side necks was fitted with a fritted glass gas dispersion tube reaching to the bottom center of the flask and equipped with a stopcock. The center neck was fitted with an adapter tube to carry the carbon dioxide past a manometer and into a thermostated gas buret. The other side neck was fitted with a device, Fig. 1, built with the help of John A. Brown, arranged to permit a weighed capsule of bromine to be admitted and broken at the proper time.

Table 1

Effect of Standing upon Accuracy of Titration

Run 38, DBPHBA in 80.2 per cent Acetic Acid, 20.0°

 $(\text{DBPHBA})_0 = 0.005595 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0182 \text{ N}$ $(\text{Br}_2)_0 = 0.007498 \text{ M}$ $(\text{LiBr}) = 0.0951 \text{ M}$

25.0 ml. of reaction mixture per flask

Elapsed time (min.)	Titer (ml.)	Titration delay (min.)	Rate constant	Per cent reaction
10.2	19.57	20	0.937	7.4
12.1	19.32	0	0.999	9.2
38.6	17.45	30	0.875	22.7
41.2	17.23	0	0.894	24.2
73.6	15.52	25*	0.867	36.5
75.2	15.48	0	0.857	36.8
156.8	12.77	0	0.828	56.3
242.0	11.11	0	0.813	68.1
408.1	9.31	0	0.797	81.2
557.2	8.36	0	0.812	88.0
843.0	6.28	0		103

*This sample was allowed to stand under a nitrogen atmosphere.

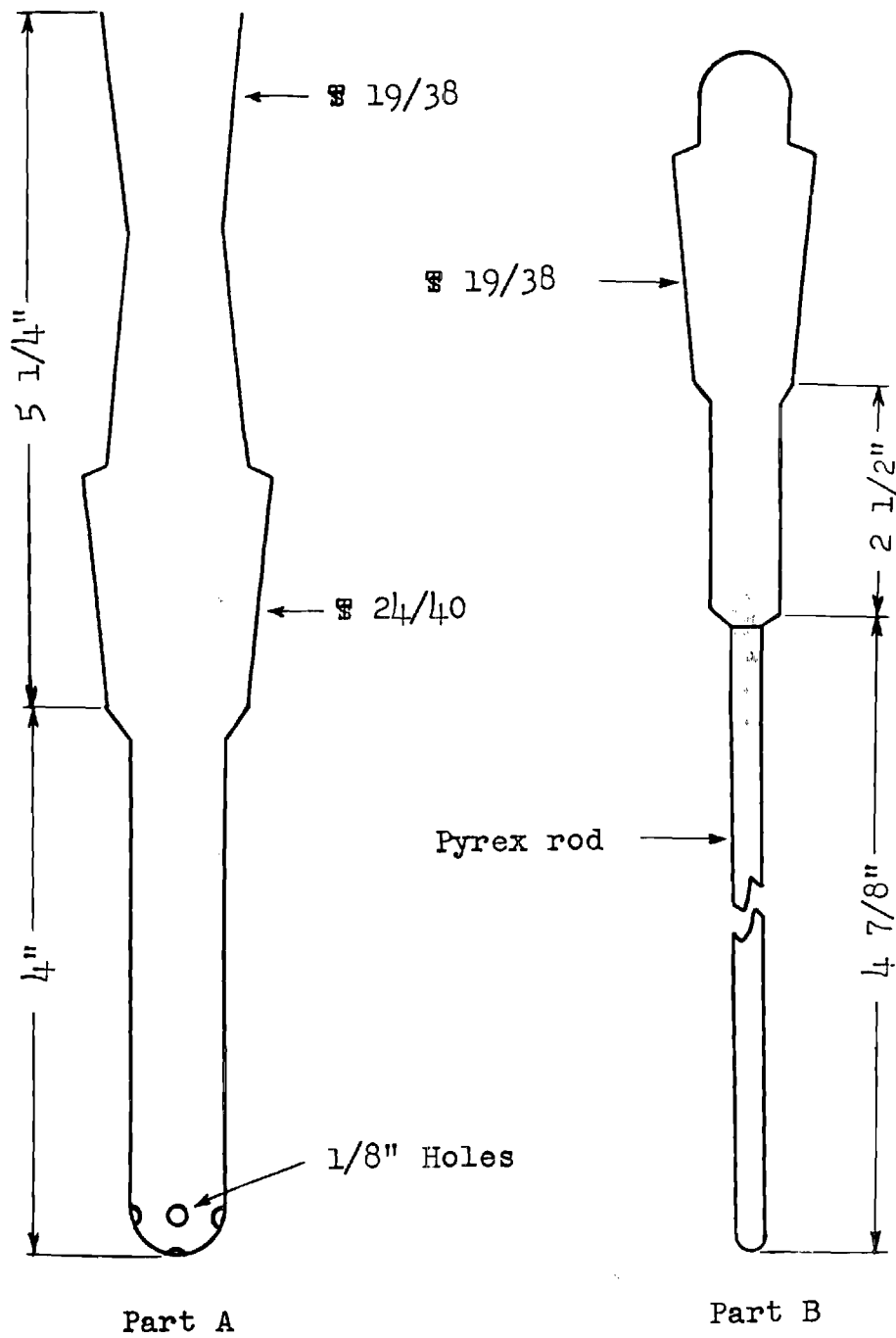


Fig. 1

The reaction flask was clamped to a ring stand and placed in the thermostat so that it was about half immersed in the water. The flask was shaken (75) by means of a "Mix-Master" motor clamped to the same ring stand. The motor was equipped with a 140 gram weight mounted eccentrically on one of its shafts. By varying the speed of the motor and the distance between the flask and the motor, the frequency and magnitude of the shaking motion could be varied. Extremely fast and vigorous shaking could be obtained in this manner. It was necessary to bolt the ring stand to the desk top to prevent it from "walking" while the shaker was in operation.

In conducting a run with this equipment, 300 ml. of DBPHBA solution is placed in the flask and carbon dioxide passed through the system for 45 minutes while the shaker is in operation. The carbon dioxide is turned off and shaking continued until the manometer level indicated that equilibrium has been obtained and no more gas is coming out of solution. The shaker is then turned off, Part B (as shown) is withdrawn from Part A, a thin walled glass capsule containing a weighed quantity of bromine¹ is dropped into Part A (which is in a side neck of the flask) and Part B is inserted forcibly to crush the bromine capsule and at

75. c.f. E. J. Corey, J. Am. Chem. Soc., 74, 5897 (1952).

¹ It was shown experimentally that the concentration of bromine could be accurately determined by weighing.

the same time to close the opening in the neck. The manometer is leveled by manipulating the mercury level in the buret, the shaker is turned on, and the time noted. The volume of gas displaced at given time intervals is then determined by means of the buret.

CHAPTER V

PRELIMINARY EXPERIMENTS IN AQUEOUS DIOXANE AND ETHANOL

The first kinetic run made was on the rate of bromination of DBSA in aqueous dioxane at 25.0°C . The aqueous dioxane solution was prepared by mixing 40.0 ml. of dioxane with 50 ml. of a water solution containing 0.004 moles of sodium chloroacetate and 0.001 moles of chloroacetic acid. The reaction was started by adding 10.0 ml. of saturated bromine water to a solution of 0.296 grams (0.001 moles) of DBSA in the above mentioned solvent. Ten milliliters of bromine water was also added to a like volume of the identical solvent. The reaction was followed by the withdrawal of 10.0 ml. portions as previously described. The results are shown in Table 2.

The formation of a precipitate was noted in the reaction flask during the initial addition of bromine. After all of the aliquots had been withdrawn, 0.20 grams of a light yellow solid melting at $135-136^{\circ}$ (with decomposition) was recovered by filtration of the contents of the reaction flask. This compound was identified as tribromophenol bromide (TBPB) by the following data:

1. The observed melting point was reasonably close

Table 2

Run 1, DBSA in Aqueous Dioxane, 25.0°.

 $(\text{DBSA})_0 = 0.010 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0204 \text{ N}$ $(\text{Br}_2)_0 = 0.0170 \text{ M}$

10.0 ml. aliquots withdrawn

Time elapsed (sec.)	Titer (ml.)	Per cent reaction
blank	16.66	0
75	8.25	86
155	2.40	146
261	2.16	148
588	1.70	153
1213	2.10	149
4054	1.40	156
4896	3.75	132

to that reported by Irvine and Smith (76) of 140-141° with decomposition. The melting point has been reported as low as 118° (11) apparently because of the extreme instability of the compound and the resulting difficulty of purification.

2. Tribromophenol was substituted for DBSA and reacted with bromine under conditions identical with those previously given and a similar rate of bromine consumption was observed and a yellow solid melting at about 135° was obtained. The results are shown in Table 3.

3. The bromine content of this substance was determined according to the method of Umhoeffer (77) and titration with standard silver nitrate and 0.1 per cent eosin as an indicator. The substance was found to be 77.2 per cent bromine (calculated bromine for TBPB is 78.02 per cent).

4. The tests reported by Lloyd (78) based on the reaction of TBPB with organic bases such as aniline, ben-zidine, etc. gave positive results.

Since TBP was not the final product of the reaction

76. F. M. Irvine and J. C. Smith, J. Chem. Soc., 1927, 75.

77. R. R. Umhoeffer, Ind. & Eng. Chem., Anal. Ed., 15, 383 (1943).

78. S. J. Lloyd, J. Am. Chem. Soc., 27, 7 (1905).

Table 3

Run 2, TBP in Aqueous Dioxane, 25.0°.

 $(\text{TBP})_0 = 0.0100 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0204 \text{ N}$ $(\text{Br}_2)_0 = 0.0181 \text{ M}$

10.0 ml. aliquots withdrawn

Time elapsed (min.)	Titer (ml.)	Per cent reaction
1.9 (blank)	16.60	
4.4	14.45	33
7.25 (blank)	16.40	
7.6	10.45	73
18.4 (blank)	16.28	
18.75	8.45	94
31.9 (blank)	16.12	
33.0	(8.80)	90
87.75 (blank)	16.10	
88.0	7.25	(106)

The anomalies observed must be attributed to the heterogeneous nature of the reaction, i.e., TBPB precipitates out and occasionally was withdrawn with the aliquots; and to the observed loss of bromine from the solvent, either by volatilization or reaction with the solvent.

in the kinetic run, it is understandable why more than 100 per cent reaction was observed. Tribromophenol was apparently formed almost instantly and in turn reacted at a finite rate with the remaining bromine in the solution to form TBPB. The several anomalies appearing in the table are explained by the fact that some of the precipitated TBPB was withdrawn with an aliquot. The fourth bromine atom on TBPB is labile enough to be at least partially titratable.

It was found by experiment that if the dioxane-water-DBSA-Bromine mixture was made about 0.1 N in perchloric acid, no TBPB would be precipitated from the mixture. Run 3 was made under these conditions. (Table 4). The DBSA solution was prepared by dissolving 0.2955 grams of DBSA (0.001 mole) in 40.0 ml. of dioxane, 40.0 ml. of distilled water and 0.85 ml. of 70 per cent perchloric acid (to give a pH of 1). A blank solution was also prepared in the same manner. The bromine solution was prepared by diluting 50 ml. of saturated bromine water with 25 ml. of distilled water (to prevent the accidental inclusion of any undissolved bromine). At recorded times, 20.0 ml. of the bromine solution was added to both the DBSA solution and the blank and the rate of the reaction was followed by the sample withdrawal technique. No precipitate was observed in the reaction flask until the un-

Table 4

Run 3, DBSA in Aqueous Dioxane, 25.0°.

 $(\text{DBSA})_0 = 0.010 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0207 \text{ N}$ $(\text{Br}_2)_0 = 0.0196 \text{ M}$ $(\text{HClO}_4) = 0.1 \text{ M}$

10.0 ml. aliquots withdrawn

Time elapsed (sec.)	Titer (ml.)	Per cent reaction
61.5	7.40	119
134	7.25	121
354	7.24	121
962	7.12	122
2316	7.02	123

used reaction mixture was diluted with water, whereupon TBP precipitated in 50 per cent yield. This should be compared with the fact that using the chloroacetic acid buffer (pH of 3.4), TBPB was recovered from the reaction mixture. The reason for obtaining greater than 100 per cent consumption of bromine in this case is not clear. It may be due to an inaccurate determination of the initial bromine concentration, a loss of bromine from volatilization or reaction with the solvent, or the partial formation of TBPB. It appears that the formation of TBPB is definitely retarded by strong acids.

Run 5 was made in the identical dioxane-water-perchloric acid solvent at -0.8°C . Even at this low temperature, 100 per cent reaction was observed in the first few minutes followed by a slow drop in the titer. It was impossible to slow the reaction by cooling because it was found that the mixture was frozen at -7°C . Furthermore, dioxane was found to form an insoluble addition compound with bromine. It was therefore decided that dioxane was an unsatisfactory solvent for the study of the kinetics of the brominative decarboxylation.

A solvent was prepared from 40.0 ml. of ethanol, 40.0 ml. of distilled water, and 0.85 ml. of 70 per cent perchloric acid. This solvent was used to dissolve 0.30 grams (0.001 mole) of DBSA. A bromine solution was pre-

pared from 50.0 ml. of saturated bromine water and 50.0 ml. of ethanol. A kinetic run was then made at -6°C . The run was started by adding 20.0 ml. of the bromine solution to the DBSA solution and to a like volume of the solvent. The rate was followed by the withdrawal of 10 ml. aliquots as previously described. The data observed is recorded in Table 5.

The data from this run is only of qualitative value. The electric stop watch was inadvertently stopped for a very short period of an unknown number of seconds between the fourth and fifth points, hence an unknown error was introduced in the rate constants. It is also obvious from the blanks that bromine has reacted with the solvent. This factor introduces further error in the rate constants. No more runs were made in this solvent.

Table 5

Run 5, DBSA in Aqueous Ethanol, -6° . $(\text{DBSA})_0 = 0.010 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0246 \text{ N}$ $(\text{Br}_2)_0 = 0.0055 \text{ M}^1$ $(\text{HClO}_4) = 0.1 \text{ M}$

10.0 ml. aliquots withdrawn

Time elapsed (min.)	Titer (ml.)	Rate constant l./mol.min.	Per cent reaction
1.2 (blank)	4.48		
1.3	3.41	22.1	24
5.8 (blank)	4.41		
5.5	2.42	17.9	46
12.0 (blank)	4.40		
10.3	1.32	8.8	71
23.1 (blank)	4.36		
21.3	0.60	16.6	87
27.4 (blank)	4.26		
25.8	0.24	(18.8)	95
58.0 (blank)	4.23		
58.1	0.08	(12.4)	98

¹This concentration was obtained by extrapolating the value of the blank titer to zero time.

CHAPTER VI

KINETIC MEASUREMENTS IN AQUEOUS ACETIC ACID

When dioxane and ethanol proved to be unsatisfactory solvents for the study of the brominative decarboxylation, aqueous acetic acid seemed to be a reasonable next choice. Because the compounds were soluble in acetic acid and since tribromophenol is formed very slowly from DBSA and DBPHBA in glacial acetic acid, it was felt that a suitable acetic acid-water mixture¹ could be found in which the reaction would proceed at a reasonable rate.

3,5-Dibromo-2-hydroxybenzoic Acid (DBSA).--A 50-50 (volume per cent) mixture of acetic acid and water, 0.1 N in perchloric acid, was prepared and used as the solvent in Runs 6 and 7 at -8°C . The rate was followed by withdrawing 10.0 ml. portions of the reaction mixture. The data obtained from Run 7 is shown in Table 6. The reaction was rather fast for accurate results and suggested that a less aqueous solvent might be more desirable from a kinetic viewpoint.

¹Unless otherwise stated, all solvents are referred to by the per cent of acetic acid by weight. In Runs 7-17 the solvents are made up by volume per cent, with no correction being made for any water in the glacial acetic acid or perchloric acid.

Table 6

Run 7, DBSA in 50 per cent (by Vol.) Acetic Acid, -8° . $(\text{DBSA})_0 = 0.00500 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0246 \text{ N}$ $(\text{Br}_2)_0 = 0.00576 \text{ M}$ $(\text{HClO}_4) = 0.1 \text{ M}$

10.0 ml. aliquots withdrawn

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
blank*	4.67*		0
1.5	3.77	37.4	22
2.9	3.26	31.4	35
4.3	2.95	29.6	43
5.7	2.73	27.1	48
11.6	2.11	23.2	63
16.5	1.80	22.4	71
23.8	1.45	23.0	80
34.2	1.14	24.8	87
52.5	0.86	(50.5)	98

*Where only one blank titer is given, that is the average of two or more blanks taken during the course of a reaction in which no significant change in blank titer was noted.

Further experiments indicated that the rate of the reaction of DBSA with bromine could be followed satisfactorily in 75 per cent acetic acid at 20°. A solution of DBSA was prepared by dissolving 0.1480 grams of DBSA in 80.0 ml. of solvent prepared by mixing 150.0 ml. of glacial acetic acid, 50.0 ml. of distilled water, and 1.70 ml. of 70 per cent perchloric acid. A bromine solution was prepared by mixing 75.0 ml. of glacial acetic acid, 15.0 ml. of saturated bromine water, and 10.0 ml. of distilled water. Run 10 was started by adding 20.0 ml. of the bromine solution to the DBSA. The initial concentration of the bromine was determined from a blank solution. The rate was followed by the withdrawal of 15.0 ml. aliquots. The data obtained from Run 10 are shown in Table 7.

The sample withdrawal technique of following this reaction had the inherent disadvantages that it was difficult to know exactly when the reaction in any aliquot portion was stopped and that with the withdrawal of each aliquot, the volume of air above the liquid was increased and thus bromine may have escaped from the liquid phase. For these reasons it seemed desirable to change to the technique which has been described in Chapter IV, whereby for any one run, a number of reactions were run concurrently in separate flasks.

It also seemed desirable to prepare the solvent with

Table 7

Run 10, DBSA in 75 per cent (by Vol.) Acetic Acid, 20.0°.

(DBSA)₀ = 0.00500 M

(Na₂S₂O₃) = 0.0247 N

(Br₂)₀ = 0.00760 M

(HClO₄) = 0.1 M

15.0 ml. aliquots withdrawn

Time elapsed (min.)	Titer (ml.)	Rate constant. (l./mol.min.)	Per cent reaction
1.9	9.00	2.80	3.8
9.0	8.20	2.95	17.0
19.0	7.30	2.98	31.8
32.1	6.50	2.96	45.0
55.9	5.52	2.98	61.2
85.7	4.73	3.08	74.2

greater precision and with a more reproducible technique so that the exact concentrations of the components could be determined. Run 12 shows the results of this first attempt at standardization. The solvent was prepared by mixing 375.0 ml. of glacial acetic acid-(a), 125 ml. of distilled water, and 4.25 ml. of 70 per cent perchloric acid. The acetic acid and water were measured at 20° and the mixture was thermostated at 20°. After mixing and equilibrating, the total volume of the mixture was found to be 487.3 ml. This solvent was then used for the preparation of the DBSA solution and the blanks. The brominating solution was prepared by mixing 20.0 ml. of bromine water and 60.0 ml. of glacial acetic acid-(a). The DBSA solution was prepared by dissolving 0.6133 grams of DBSA in 325.0 ml. of the solvent. Twenty milliliters of the DBSA solution was placed in each of eight 50 ml. volumetric flasks. Twenty milliliters of the solvent was also placed in each of three 50 ml. volumetric flasks to be used as blanks. The reactions were started by adding 5.0 ml. of the bromine solution to each flask at a recorded time. The rate was then followed by the technique that has been previously described. The results of Run 12 are recorded in Table 8.

It is apparent from the blanks that were run in Run 12, that there was a pronounced side reaction in the solvent. This side reaction does not significantly affect

Table 8

Run 12, DBSA in 75 per cent (by Vol.) Acetic Acid-(a), 20.00.

(DBSA)₀ = 0.00510 M

(Na₂S₂O₃) = 0.0230 N

(Br₂)₀ = 0.00630 M*

(HClO₄) = 0.0822 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
5.8 (blank)	13.63		
14.7	11.20	3.05	23.5
28.3	9.64	3.12	38.2
45.0	8.87	2.54	45.3
62.1	7.38	3.08	59.4
63.3 (blank)	13.07		
89.9	6.25	3.06	69.8
120.1	5.40	3.11	77.8
150.2	4.64	3.47	84.9
180.0	4.06	3.78	90.5
180.8 (blank)	11.82		
0.0 (blank)**	13.42		

*Assuming a reaction of the bromine with the solvent, the initial bromine concentration was determined by extrapolating the bromine titers back to zero time. The rate constants were calculated on the basis of this extrapolated bromine concentration.

**This blank was a 5.0 ml. portion of stock brominating solution which was titrated for bromine after the run had been completed.

the rate constants until the principal reaction has been greatly slowed down by the depletion of the concentration of the DBSA.

Runs 13-17 were made in order to track down the source of the side reaction which had been observed. In Run 13, the source of bromine was changed, in Run 14 the concentration of bromine was halved, in Run 15 sulfuric acid was substituted for perchloric acid, in Run 16 the concentration of DBSA was doubled, and in Run 17 the concentration of DBSA was halved. In other than these respects Run 12-17 were identical. Tables 9-13 show the results of these runs.

It should be noted that Runs 12-17 all have a minimum rate constant in the middle of the reaction. Since the blanks showed a considerable fall in titer in all of the runs, it must be assumed that there is a brominatable impurity in the acetic acid. The fall in rate constant that was generally observed at the beginning of a run is probably due to the deactivation of the bromine by the bromide ions that are formed during the reaction¹. This troublesome reaction between bromine and bromide ions to form tribromide ions has been considered extensively by

¹Since the solvent was 0.10 N in perchloric acid, the small increase in (H^+) with reaction must have produced only a relatively small drop in rate constant.

Table 9

Run 13, DBSA in 75 per cent (by Vol.) Acetic Acid-(a), 20.0°.

(DBSA)₀ = 0.005209 M

(Na₂S₂O₃) = 0.0230 N

(Br₂)₀ = 0.00497 M*

(HClO₄) = 0.0822 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
2.1 (blank)	10.77		
15.0	8.63	3.20	20.1
29.9	7.30	3.21	32.4
45.0	6.25	2.99	42.2
60.2	5.54	2.87	48.7
60.4 (blank)	10.11		
90.5	4.30	3.08	60.2
120.1	3.52	3.11	67.5
150.3	2.87	3.32	73.4
180.1	2.31		98.7
180.3 (blank)	8.99		

*Calculated from values of blank titers extrapolated to zero time.

Table 10

Run 14, DBSA in 75 per cent (by Vol.) Acetic Acid-(a), 20.0°.

(DBSA)₀ = 0.005209 M

(Na₂S₂O₃) = 0.0230 N

(Br₂)₀ = 0.00242 M

(HClO₄) = 0.0822 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
1.1 (blank)	5.27		
20.0	3.54	4.17	32.6
40.1	2.67	3.73	49.2
55.1	2.20	3.65	58.3
88.3	1.77	2.94	66.4
88.2 (blank)	4.17		
98.2	1.54	3.04	70.8
113.2	1.26	3.15	76.0
128.3	0.98	3.38	81.4
143.3	0.80	3.46	84.8
143.4 (blank)	3.83		

Table 11.

Run 15, DBSA in 75 per cent (by Vol.) Acetic Acid-(a), 20.0°.

(DBSA)₀ = 0.002733 M

(Na₂S₂O₃) = 0.0230 N

(Br₂)₀ = 0.00567 M

(H₂SO₄) = 0.0544 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titler (ml.)	Rate constant (l./mol.min.)	Per cent reaction
1.0 (blank)	12.33		
10.0	11.08	4.38	20.8
20.0	10.29	4.11	34.4
30.0	9.61	4.13	45.8
45.0	8.79	4.32	59.7
50.7 (blank)	11.70		
60.0	8.21	4.39	69.2
80.0	7.57	4.82	80.2
102.0	7.03	5.60	89.4
120.0	6.70	6.70	94.8
120.0 (blank)	11.00		

Table 12

Run 16, DBSA in 75 per cent (by Vol.) Acetic Acid-(a), 20.0°.

(DBSA)₀ = 0.01014 M

(Na₂S₂O₃) = 0.0226 N

(Br₂)₀ = 0.00580 M

(HClO₄) = 0.0822 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
0.6 (blank)	12.82		
5.4	10.92	3.05	14.8
11.7	9.32	2.95	27.4
14.1	8.88	2.86	30.9
25.9	7.00	2.70	45.6
26.6 (blank)	12.40		
34.6	5.98	2.65	53.5
46.8	4.94	2.56	61.6
70.6	3.60	2.36	72.0
90.6	2.62	2.44	79.4
94.2 (blank)	11.72		

Table 13

Run 17, DBSA in 75 per cent (by Vol.) Acetic Acid-(a), 20.0°.

(DBSA)₀ = 0.00254 M

(Na₂S₂O₃) = 0.0226 N

(Br₂)₀ = 0.00481 M

(HClO₄) = 0.0822 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
4.1 (blank)	10.62		
5.4	10.09	4.13	9.9
12.6	9.56	3.95	19.3
25.0	8.72	3.89	34.3
40.1	8.00	3.87	47.3
59.1	7.29	3.99	59.9
78.3	6.68	4.47	70.5
100.0	6.15	4.66	80.0
119.2	5.72	5.27	87.9
122.2 (blank)	9.26		

Wilson and Soper (36), Bradfield, Jones, and Orton (31), and by Kuivila and Esterbrook (42) and will be dealt with at length later in this work. The general trend of the rate constants in Runs 12-17 can then be explained by the assumption that initially the dominating reaction is the bromination of the DBSA, which reaction is greatly slowed by formation of bromide ion. As the DBSA is consumed the reaction with solvent becomes more important and ultimately is proceeding at a greater rate than the DBSA reaction; this would cause an apparent increase in rate constant.

Since commercial acetic acid was not purified sufficiently by merely distilling, it was purified in the manner described in Chapter III to make acetic acid-(b). Assuming water to be the only impurity and knowing the water content of this acetic acid, its density could be obtained from the International Critical Tables (79a). From this density and the known densities of water (79b) and of 70 per cent perchloric acid (79c), a solvent 75 per cent acetic acid by weight was prepared as described in the appendix.

Runs 18 and 19 were made in this type solvent and are recorded in Tables 14 and 15. These data show a con-

79a. International Critical Tables, 1st Ed., Vol.III, p. 123.

79b. Ibid., p. 25.

79c. Ibid., p. 54.

Table 14

Run 18, DBSA in 75.0 per cent Acetic Acid-(b), 20.0°

 $(\text{DBSA})_0 = 0.005006 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0211 \text{ N}$ $(\text{Br}_2)_0 = 0.003764 \text{ M}$ $(\text{HClO}_4) = 0.100 \text{ M}$

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate Constant (l./mol.min.)	Per cent reaction
(blank)	8.92		0.0
5.1	8.20	3.40	8.1
15.4	7.30	2.80	18.1
30.9	5.98	3.00	32.9
45.8	5.14	2.95	42.4
63.2	4.52	2.76	49.3
95.5	3.50	2.74	60.8
133.9	2.81	2.59	68.5

Table 15

Run 19, DBSA in 75.0 per cent Acetic Acid-(b), 20.0°

 $(\text{DBSA})_0 = 0.005006 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0211 \text{ N}$ $(\text{Br}_2)_0 = 0.007495 \text{ M}$ $(\text{HClO}_4) = 0.100 \text{ M}$

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	17.76		0.0
5.5	16.30	3.36	12.3
15.4	14.50	3.10	27.5
30.2	12.65	2.99	43.1
45.1	11.50	2.81	52.8
60.8	10.36	2.90	62.4
84.7	9.48	2.70	69.8
105.3	8.85	2.65	75.1
135.4	8.21	2.56	80.5
165.5	7.77	2.48	84.2
186.2	7.49	2.47	86.6

sistent downward drift in the rate constants which was assumed to be due to the accumulation of bromide ion. Run 20, Table 16, confirms this assumption. In this run, 0.0624 grams (0.524 millimoles) of potassium bromide was dissolved in enough DBSA solution to make 100 ml. It is seen that the rate constant was depressed but not as much as one might expect when one compares the amount of bromide ion added with the amount that is liberated in the reaction. It was suspected that the addition of potassium bromide caused a positive salt effect which partially masked the negative effect of the bromide ion. This was demonstrated to be true in Run 21, Table 17, in which potassium perchlorate was added to the DBSA solution. The data shows a slight increase in the rate constant but nothing really striking.

The decrease in rate by bromide ion is shown most markedly in Run 28, Table 18. The solvent in this run was made 0.1 N in lithium bromide (as described in the appendix) and was calculated to be 75.2 per cent acetic acid by weight.

In order to determine the quantitative effect of bromide ion on the rate of the reaction of bromine with DBSA, a series of runs was made in 75 per cent acetic acid in which the hydrogen ion concentration was held constant at 0.300 M, the ionic strength was held constant at 0.3 and

Table 16

Run 20, DBSA in 75.0 per cent Acetic Acid-(b), 20.0°

(DBSA)₀ = 0.005006 M

(Na₂S₂O₃) = 0.0211 N

(Br₂)₀ = 0.007436 M

(KBr) = 0.004195 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	17.62		0.0
7.2	16.06	2.79	13.2
15.8	14.78	2.55	24.0
23.6	13.85	2.48	31.8
33.7	12.72	2.53	41.3

Table 17

Run 21, DBSA in 75.0 per cent Acetic Acid-(b), 20.0°

(DBSA)₀ = 0.005079 M

(Na₂S₂O₃) = 0.0211 N

(Br₂)₀ = 0.004857 M

(KClO₄) = 0.004013 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	11.51		0.0
5.7	10.41	3.33	9.6
15.3	9.14	3.11	20.6
25.8	8.11	3.14	29.5
38.9	7.10	3.08	38.3

Table 18

Run 28, DBSA in 75.2 per cent Acetic Acid-(b), 20.0°

 $(\text{DBSA})_0 = 0.005032 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0223 \text{ N}$ $(\text{Br}_2)_0 = 0.003925 \text{ M}$ $(\text{LiBr}) = 0.1024 \text{ M}$ $(\text{HClO}_4) = 0.100 \text{ M}$

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	8.80		0.0
330	6.00	0.268	31.8
645	4.61	0.255	47.6
1260	3.13	0.241	64.4
1665	2.56	0.233	70.9
2076	2.11	0.230	76.0
2796	1.55	0.229	82.4

the bromide ion concentration was varied between zero and 0.300 M. The results of these runs are shown in Tables 19-25.

In a similar manner, to determine the effect of hydrogen ion concentration on the rate, a series of runs was made in which the bromide ion concentration was held constant at 0.300 M, the ionic strength was held constant at 0.3, and the hydrogen ion concentration was varied between 0.100 and 0.300 M. The results of these runs are shown in Tables 22, 26, and 27.

To demonstrate that the reaction was indeed proceeding as supposed, a 23 per cent yield (0.03 g.) of tribromophenol was isolated from 60 ml. of solution that was initially 0.0065 M in DBSA and 0.0066 M in bromine and in the solvent of Run 77, Table 27. The mixture had been allowed to stand for four days at ca. 40°. The melting point of the crude product was 88°. The low yield was due primarily to mechanical losses in the small amount of material handled.

The final series of runs on DBSA was made in order to determine the effect of the concentration of water in the solvent. Runs were made in 70, 75, and 80 per cent acetic acid, all 0.300 M in hydrogen bromide. The marked effect of the variation in solvent is shown in Tables 22, 28, and 29. A change in concentration of water by 10 per cent causes almost a 50 fold change in rate!

Table 19

Run 79, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°

(DBSA)₀ = 0.004157 M

(Na₂S₂O₃) = 0.02138 N

(Br₂)₀ = 0.007996 M

(HBr)₀ = 0.0500 M

(HClO₄) = 0.250 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.40		0.0
51.5	35.85	0.205	8.0
111.5	34.38	0.197	15.5
244	32.80	(0.148)	23.6
418	29.30	0.184	41.7
574	27.50	0.184	50.9
1158	23.58	0.175	71.1

Rate constant extrapolated to zero time (k_{app}) = 0.215

Table 20

Run 73, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°

(DBSA)₀ = 0.004313 M

(Na₂S₂O₃) = 0.02138 N

(Br₂)₀ = 0.007973 M

(HBr)₀ = 0.100 M

(HClO₄) = 0.200 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.29		0.0
61.5	36.35	(0.0988)	4.7
180	35.15	0.0804	10.6
481	32.40	0.0779	24.3
1134	28.45	0.0738	43.8
1907	25.42	0.0723	58.8
8351	18.18	0.0728	94.7

Rate constant extrapolated to zero time (k_{app}) = 0.0825

Table 21

Run 74, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°
 (DBSA)_o = 0.004196 M (Na₂S₂O₃) = 0.02138 N
 (Br₂)_o = 0.007945 M
 (HBr)_o = 0.200 M
 (HClO₄) = 0.100 M
 50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Percent reaction
(blank)	37.16		0.0
332	35.80	(0.0277)	6.9
903	33.89	0.0266	16.7
1646	31.80	0.0264	27.3
2331	30.19	0.0264	35.5
3283	28.39	0.0262	44.7
4368	26.67	0.0265	53.5

Average value of rate constant (k_{app}) = 0.0264 ± 0.4%

Table 22

Run 75, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°(DBSA)₀ = 0.004313 M(Na₂S₂O₃) = 0.02138 N(Br₂)₀ = 0.008368 M(HBr)₀ = 0.300 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	39.14		0.0
776	37.32	0.0148	9.0
1460	36.07	0.0141	15.2
2812	33.76	0.0142	26.7
4118	32.08	0.0139	35.0
6266	(30.30)	(0.0126)	43.8
7709	28.38	0.0141	53.3

Average value of rate constant (k_{app}) = 0.0142 ± 1.5%

Table 23

Run 81, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°

(DBSA)_o = 0.004317 M

(Na₂S₂O₃) = 0.02138 N

(Br₂)_o = 0.007821 M

(HBr)_o = 0.0100 M

(HClO₄) = 0.290 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	36.58		0.0
20.5	34.62	0.655	9.7
35.6	33.56	0.611	15.0
68.9	31.12	0.638	27.1
131.2	27.61	0.667	44.4
225.3	24.71	0.626	58.8
363.2	22.01	0.606	72.2

Average value of rate constant (k_{app}) = $0.634 \pm 3.0\%$

Table 24

Run 83, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°

(DBSA)₀ = 0.004349 M

(Na₂S₂O₃) = 0.02135 N

(Br₂)₀ = 0.007932 M

(HBr)₀ = 0.0100 M

(HClO₄) = 0.290 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.15		0.0
20.5	34.87	0.758	11.2
37.1	33.43	0.727	18.3
62.8	31.58	0.700	27.4
117.2	28.58	0.678	42.2
191.6	25.80	0.658	55.8
263.8	24.00	0.638	64.7

Rate constant extrapolated to zero time (k_{app}) = 0.830

Table 25

Run 82, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°
 (DBSA)₀ = 0.004499 M (Na₂S₂O₃) = 0.02135 N
 (Br₂)₀ = 0.007942 M
 (HClO₄) = 0.300 M
 50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.20		0.0
5.0	35.95	(1.58)	5.9
14.5	34.21	1.39	14.2
30.0	32.00	1.29	24.7
55.0	29.01	1.29	38.9
90.0	26.45	1.20	51.0
148.3	23.30	1.19	65.9

Rate constant extrapolated to zero time (k_{app}) = 1.65

Table 26

Run 76, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°

(DBSA)₀ = 0.004238 M

(Na₂S₂O₃) = 0.02138 N

(Br₂)₀ = 0.007964 M

(HBr)₀ = 0.200 M

(LiBr) = 0.100 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.25		0.0
354	35.72	(0.0292)	7.7
1037	33.65	0.0255	18.2
1665	31.90	0.0257	27.0
2917	29.18	0.0256	40.7
3825	27.51	0.0262	49.1
5271	25.68	0.0257	58.4

Average value of rate constant (k_{app}) = 0.0257 \pm 0.6%

Table 27

Run 77, DBSA_a in 75.0 per cent Acetic Acid-(b), 20.0°

(DBSA)_o = 0.004060 M (Na₂S₂O₃) = 0.02138 N

(Br₂)_o = 0.007898 M

(HBr)_o = 0.100 M

(LiBr) = 0.200 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	36.94		0.0
194	35.00	0.0726	10.2
928	30.00	0.0694	36.6
1362	28.10	0.0676	46.6
1765	26.50	0.0688	55.0
2375	24.81	0.0681	63.9
3225	23.15	0.0670	72.6

Average value of rate constant (k_{app}) = 0.0689 \pm 2.0%

Table 28

Run 80, DBSA_a in 70.0 per cent Acetic Acid-(b), 20.0°

(DBSA)₀ = 0.004388 M

(Na₂S₂O₃) = 0.02138 N

(Br₂)₀ = 0.008120 M

(HBr)₀ = 0.300 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.98		0.0
190	35.85	0.0733	10.4
375	34.15	0.0715	18.6
980	29.96	0.0707	39.1
1437	27.79	0.0697	49.6
2805	23.65	0.0692	69.8
15329	17.57	-	99.6

Average value of rate constant (k_{app}) = 0.0709 \pm 1.7%

Table 29

Run 78, DBSA_a in 80.0 per cent Acetic Acid-(b), 20.0°
 (DBSA)_o = 0.004345 M (Na₂S₂O₃) = 0.02138 N
 (Br₂)_o = 0.007866 M
 (HBr)_o = 0.300 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	36.79		0.0
1327	36.28	(0.00243)	2.5
4715	35.52	0.00176	6.3
10652	34.08	0.00178	13.3
20036	32.25	0.00174	22.6
30110	30.50	0.00174	31.2
39456	29.00	0.00178	38.5

Average value of rate constant (k_{app}) = 0.00176 \pm 0.9%

3,5-Dibromo-4-hydroxybenzoic acid (DBPHBA).--The first study made on DBPHBA was in 75 per cent acetic acid and the data obtained are recorded in Table 30. It is seen that in this solvent, the reaction is extremely rapid, in fact too rapid for accurate measurement by the technique being used. The next run was in 80 per cent acetic acid, Table 31, and while the rate of the reaction had been slowed considerably, it was still too fast for accurate measurement.

It was discovered that a very convenient rate was obtained if the reaction mixture was made 0.1 M in bromide ion. This was done by preparing a lithium bromide solution in water and using this in preparation of the solvent (see appendix). Runs 26-47 were run in solvent that was actually 80.2 per cent acetic acid by weight.

Some of the results obtained in this solvent are shown in Tables 32-35. Run 35 using freshly prepared DBPHBA was made to determine whether freshly prepared DBPHBA reacted at the same rate as DBPHBA that was 7 months old. The results shown in Table 36 were very inconclusive in this respect. The DBPHBA used through Run 35 had only been purified by recrystallization from acetic acid. The effect of further recrystallization from acetonitrile was tested in Run 36. The results shown in Table 37 are inconclusive. Run 37, Table 38, was made primarily for the

Table 30

Run 22, DBPHBA in 75.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.005004 M

(Na₂S₂O₃) = 0.0211 N

(Br₂)₀ = 0.005794 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	13.73		0.0
1.1	9.15	94.8	22.0
2.0	7.10	100.9	55.9
3.1	5.88	96.6	66.2
4.75	4.70	96.3	76.1
6.3	4.07	94.3	81.5
7.65	3.65	94.8	85.0
8.7	3.48	91.0	86.4
13.0	2.87	88.6	91.6

Table 31

Run 23, DBPHBA in 80.0 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.005012 M (Na₂S₂O₃) = 0.0211 N
 (Br₂)₀ = 0.005815 M
 (HClO₄) = 0.100 M
 25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	13.78		0.0
1.1	11.58	35.0	18.5
2.6	9.60	32.5	35.5
4.7	8.03	32.2	48.5
8.7	6.25	30.7	63.4
11.4	5.61	29.0	68.8
15.6	4.90	27.4	74.8

Table 32

Run 26, DBPHBA in 80.2 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.004974 M

(Na₂S₂O₃) = 0.0223 N

(Br₂)₀ = 0.003822 M

(LiBr) = 0.1032 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	8.57		0.0
17.0	8.12	(0.648)	5.3
45.3	7.42	0.675	13.4
74.0	6.78	0.697	20.9
115.0	6.05	0.696	29.4
176.1	5.20	0.689	39.3
236.5	4.58	0.675	46.6

Table 33

Run 27, DBPHBA in 80.2 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.004974 M

(Na₂S₂O₃) = 0.0223 N

(Br₂)₀ = 0.004170 M

(LiBr) = 0.1032 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	9.35		0.0
126	6.40	0.710	31.6
210	5.25	0.708	43.9
345	4.12	0.673	55.9
522	3.12	0.667	66.6
1110	1.63	0.637	82.6

Table 34

Run 29, DBPHBA in 80.2 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.005053 M

(Na₂S₂O₃) = 0.0223 N

(Br₂)₀ = 0.008117 M

(LiBr) = 0.1024 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	18.20		0.0
78	14.38	0.735	33.7
138	12.82	0.695	47.5
198	11.60	0.697	58.3
324	10.14	0.663	71.1
546	8.69	0.652	83.9
1386	7.25	(0.583)	96.6

Table 35

Run 34, DBPHBA in 80.2 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.006354 M

(Na₂S₂O₃) = 0.0200 N

(Br₂)₀ = 0.005900 M

(LiBr) = 0.0998 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	14.75		0.0
11	13.80	(0.982)	6.4
29	12.60	0.932	14.6
55	11.20	0.899	24.1
98	9.54	0.861	39.5
122	8.82	0.849	44.9
208	6.94	0.820	59.1
279	5.93	0.799	66.8
364	5.05	0.780	73.5
453	4.31	0.776	79.1
1050	2.20	(0.718)	95.0

Table 36

Run 35, DBPHBA in 80.2 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.006691 M

(Na₂S₂O₃) = 0.0200 N

(Br₂)₀ = 0.005568 M

(LiBr) = 0.0951 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	13.92		0.0
10.6	13.10	0.886	5.9
24.7	12.17	0.861	12.6
48.9	10.87	0.839	21.9
73.4	9.72	0.850	30.2
109.1	8.53	0.823	38.7
148.6	7.55	0.794	45.8
194.2	6.55	0.794	52.9
244.9	5.75	0.778	58.7
313.5	4.97	0.751	64.3
461.2	3.68	0.740	73.6
544.7	3.18	0.734	77.2

Table 37

Run 36, DBPHBA in 80.2 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.006674 M

(Na₂S₂O₃) = 0.0201 N

(Br₂)₀ = 0.005395 M

(LiBr) = 0.0951 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	13.42		0.0
11.0	12.55	0.938	6.5
30.3	11.39	0.866	15.1
46.2	10.60	0.842	21.0
65.2	9.81	0.818	26.9
96.6	8.68	0.806	35.3
127.3	7.78	0.800	42.0
170.3	6.80	0.786	49.3
222.4	5.87	0.775	56.3
273.5	5.15	0.767	61.6
331.2	4.50	0.760	66.5
396.0	3.93	0.751	70.7
542.3	2.90	0.761	78.4

Table 38

Run 37, DBPHBA in 80.2 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.005429 M

(Na₂S₂O₃) = 0.0201 N

(Br₂)₀ = 0.008060 M

(LiBr) = 0.0951 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	20.05		0.0
10.4	19.00	1.006	7.8
27.2	17.71	0.925	17.3
56.6	16.35	0.780	27.4
81.5	14.95	0.843	37.8
121.6	13.62	0.815	47.6
164.4	12.47	0.807	56.1
209.9	11.55	0.799	62.9
258.3	10.80	0.789	68.5
317.5	10.22	0.751	72.8
352.1	9.80	0.763	75.9
405.7	9.32	0.765	79.4
425.3	6.50		100.4

purpose of obtaining the infinity titer as an index to the purity of the DBPHBA. The 0.4 per cent excess reaction indicated is probably due in the main to the accumulation of experimental errors. The indication is that the DBPHBA is essentially pure.

Run 39, Table 39, was made to test the result of purifying DBPHBA by Method A. Although a commercial analysis indicated a trace of impurity in DBPHBA_a, this would be hard to confirm by the kinetic data. There is possibly evidence of a slight amount of non-brominatable substance (from the infinity titer), although the slight deviation from 100 is probably within the experimental error.

Table 40 shows the kinetic data obtained from DBPHBA purified by Method B. Kinetically, very little difference is noted from DBPHBA_a, however the results of a commercial analysis for bromine were more satisfactory, so DBPHBA_b was used for the remainder of the runs.

It would have been extremely useful to measure the rate of carbon dioxide evolution from DBPHBA in the reaction and to compare it with the rate of bromine consumption. Apparatus was designed and built to do this and the results are shown in Table 41. Unfortunately, the results are very inconclusive. The rate constants measured are slightly lower in magnitude and unlike the bromine consumption constants, they go through a maximum. This may have been due

Table 39

Run 39, DBPHBA_a in 80.2 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.005657 M (Na₂S₂O₃) = 0.0200 N
 (Br₂)₀ = 0.007980 M
 (LiBr) = 0.1037 M
 (HClO₄) = 0.100 M
 25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	19.95		0.0
15.6	18.46	0.923	10.5
77.5	14.85	0.844	36.1
245	10.70	0.770	65.4
384	9.12	0.749	76.6
1217	6.58	0.636	94.5
47.5 hrs.	6.00		98.6
143.2 "	5.88		99.5

Table 40

Run 43, DBPHBA_p in 80.2 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)_o = 0.004540 M (Na₂S₂O₃) = 0.0202 N
 (Br₂)_o = 0.008565 M
 (LiBr) = 0.1041 M
 (HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	21.20		0.0
16.5	20.00	0.820	10.7
48.0	18.20	0.817	26.7
93.5	16.82	0.697	39.0
141.5	15.38	0.717	51.8
190.0	14.40	0.709	60.5
248.0	13.18	0.777	71.4
299.0	12.90	0.702	73.9
496.0	11.58	0.668	85.6
1225	10.20		97.9
69.5 hrs.	10.00		99.7

Table 41

Run 44, DBPHBA₀ in 80.2 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.01176 M
 (Br₂)₀ = 0.01803 M
 (LiBr) = 0.1041 M
 (HClO₄) = 0.100 M

CO₂ displaced from 300 ml. of reaction mixture

Time elapsed (min.)	Volume of CO ₂ at STP (ml.)	Rate constant (L./mol.min.)	Per cent reaction
19	12.0	0.507	15.2
37	21.2	0.517	26.8
54	30.1	0.573	38.1
77	38.2	0.584	48.3
105	45.2	0.580	57.2
139	51.5	0.575	65.1
163	55.7	0.591	70.5
214	60.5	0.565	76.5
271	64.7	0.555	81.8
1380	76.5		97.0
1380			99.0*
46 hrs.			99.2*

*Per cent reaction measured by titrating a portion of the reaction mixture for bromine.

to a tendency of carbon dioxide to supersaturate the solution.

Run 47, Table 42, was run to determine something of the effect of hydrogen ion concentration on the rate of the reaction. It was noted that the rate of the reaction was almost doubled by halving the hydrogen ion concentration.

In Runs 48-53, Tables 43-48, hydrogen bromide was used in the solvent to supply hydrogen ion and bromide ion in place of using lithium bromide and perchloric acid. While Run 51 was made in a solvent computed to be 79.8 per cent acetic acid, the average value of the rate constant is not much greater than would have been predicted for 80.0 per cent acetic acid, although a 0.2 per cent increase in water concentration normally gave about 100 per cent increase in rate.

The most precise and valuable work on DBPHBA is recorded in Tables 49-59. This precision is due probably to two factors: 1. The runs are all made at the same ionic strength (0.3). 2. The solvents used in these runs were prepared in such a way that the concentrations of the various ingredients could be known with greater precision. In Tables 49-53, the results are shown of series of runs in which the bromide ion concentration was held constant at 0.300 M and the hydrogen ion concentration varied between 0.300 M and 0.05 M. In Tables 49, 54-57, the results

Table 42

Run 47, DBPHBA_p in 80.2 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.004606 M (Na₂S₂O₃) = 0.0201 N
 (Br₂)₀ = 0.007871 M
 (LiBr) = 0.1041 M
 (HClO₄) = 0.05005 M
 25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	19.58		0.0
9.0	18.27	1.774	11.4
15.5	17.29	1.948	20.0
31.0	15.60	1.971	34.7
50.5	14.30	1.841	46.1
78.5	12.87	1.800	58.6
105.0	11.98	1.742	66.3
133.0	11.21	1.736	73.1
180.0	10.34	1.708	80.6
240.0	9.66	1.662	86.6
315.0	9.08		91.6
1400	8.22		99.2

Table 43

Run 48, DBPHBA₀ in 80.0 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.004911 M (Na₂S₂O₃) = 0.0202 N
 (Br₂)₀ = 0.008019 M
 (HBr)₀ = 0.1072 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	19.85		0.0
17.0	19.23	(0.388)	5.1
33.5	18.51	0.449	11.0
59.5	17.69	0.434	17.8
90.0	16.77	0.442	25.3
122.0	15.98	0.439	31.8
150.0	15.38	0.436	36.8
182.5	14.65	0.448	42.8
227.0	14.00	0.435	48.1
258.5	13.48	0.442	52.4
286.5	13.20	0.431	54.7
360.0	12.58	0.407	59.8
807.0	9.81	0.416	82.6

Table 44

Run 49, DBPHBA_p in 80.0 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)_o = 0.004972 M (Na₂S₂O₃) = 0.02014 N
 (Br₂)_o = 0.008016 M
 (HBr)_o = 0.1073 M
 25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	19.90		0.0
56.0	17.80	0.441	17.0
164.0	15.15	0.428	38.5
241.5	13.95	0.413	48.2
301.0	13.10	0.418	55.1
361.5	12.45	0.415	60.4
424.5	11.70	0.434	66.4
534.0	11.02	0.419	71.9
623.5	10.68	0.396	74.7
754.0	10.02	0.403	80.0
1389	8.60	0.386	91.6
2880	7.90		97.2

Table 45

Run 51, DBPHBA₀ in 79.8 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.004798 M (Na₂S₂O₃) = 0.02199 N
 (Br₂)₀ = 0.008836 M
 (HBr)₀ = 0.1083 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	40.18		0.0
65	35.40	0.459	21.9
123	32.32	0.462	36.0
205	29.35	0.449	49.6
298	27.00	0.440	60.4
669	22.08	0.434	83.0
1025	20.56	(0.397)	90.0

Table 46

Run 50, DBPHBA₀ in 80.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.005144 M

(Na₂S₂O₃) = 0.02199 N

(Br₂)₀ = 0.008070 M

(HBr)₀ = 0.2221 M

25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	18.40		0.0
31.0	18.20	(0.0681)	1.7
94.5	18.02	0.0434	3.2
330.5	17.03	0.0485	11.7
565.0	16.19	0.0488	18.8
1132	15.11	0.0399	28.1
1380	14.72	0.0380	31.4
2055	13.51	0.0383	41.7
2783	12.62	0.0371	49.3
4792	10.58	0.0389	66.7
8348	9.01	0.0365	79.8
120 hrs. @ 40°	6.60		100.5

Table 47

Run 52, DBPHBA₀ in 80.0 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.005177 M (Na₂S₂O₃) = 0.02197 N
 (Br₂)₀ = 0.008052 M
 (HBr)₀ = 0.2221 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	36.65		0.0
321	34.42	0.0396	9.5
634	32.68	0.0382	16.8
1060	30.80	0.0365	24.8
1341	29.32	0.0387	31.1
2079	26.85	0.0379	41.6
3607	23.05	0.0383	57.7

Table 48

Run 53, DBPHBA_b in 80.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.008934 M (Na₂S₂O₃) = 0.02197 N

(Br₂)₀ = 0.005100 M

(HBr)₀ = 0.2225 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	23.21		0.0
449	20.02	0.0384	13.8
1255	15.90	0.0375	31.5
1737	14.10	0.0368	39.3
2826	10.82	0.0369	53.4
4299	8.01	0.0362	65.5
8577	3.82	0.0352	83.5

Table 49

Run 56, DBPHBA₀ in 80.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.005047 M

(Na₂S₂O₃) = 0.02195 N

(Br₂)₀ = 0.008062 M

(HBr)₀ = 0.300 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	36.73		0.0
371	35.76	0.0147	4.2
1413	33.58	0.0135	13.7
2888	30.96	0.0135	25.1
4321	29.01	0.0133	33.6
6973	26.05	0.0133	46.4
10117	23.80	0.0129	56.2

Average rate constant (k_{app}) = 0.0135 \pm 3.0%

Table 50

Run 54, DBPHBA_p in 80.0 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.004515 M (Na₂S₂O₃) = 0.02190 N
 (Br₂)₀ = 0.008252 M
 (HBr)₀ = 0.200 M
 (LiBr)₀ = 0.100 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.68		0.0
405	35.60	0.0329	10.1
1174	32.60	0.0314	24.7
1886	30.40	0.0313	35.3
2693	28.43	0.0312	44.9
4314	25.60	0.0305	58.3
8814	21.36	0.0302	79.0

Average rate constant (k_{app}) = 0.0313 \pm 1.9%

Table 51.

Run 55, DBPHBA₀ in 80.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.004458 M

(Na₂S₂O₃) = 0.02190 N

(Br₂)₀ = 0.008142 M

(HBr)₀ = 0.200 M

(KBr) = 0.100 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.18		0.0
476	34.64	0.0355	12.5
1281	31.15	0.0369	29.6
1895	29.28	0.0357	38.5
2916	26.66	0.0364	51.4
4593	23.94	0.0358	64.8
7735	21.03	0.0350	79.1

Average rate constant (k_{app}) = 0.0359 ± 1.4%

Table 52

Run 57, DBPHBA_b in 80.0 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.001499 M (Na₂S₂O₃) = 0.02195 N
 (Br₂)₀ = 0.008119 M
 (HBr)₀ = 0.100 M
 (LiBr) = 0.200 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	36.99		0.0
129	34.55	0.125	11.9
282	32.27	0.122	23.0
463	30.13	0.120	33.5
638	28.38	0.121	42.0
1296	24.25	0.117	62.1
1809	22.27	0.116	71.8

Average rate constant (k_{app}) = 0.120 \pm 2.1%

Table 53

Run 58, DBPHBA_p in 80.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.004481 M

(Na₂S₂O₃) = 0.02195 N

(Br₂)₀ = 0.008157 M

(HBr)₀ = 0.0500 M

(LiBr) = 0.250 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.16		0.0
26.0	35.50	0.408	8.1
63.5	33.40	0.416	18.4
120.5	30.69	0.429	31.7
192.5	28.27	0.422	43.6
474.0	23.10	0.397	68.9
4012	16.95		98.9

Average rate constant (k_{app}) = 0.414 ± 2.4%

Table 54

Run 63, DBPHBA₀ in 80.0 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.005128 M (Na₂S₂O₃) = 0.02185 N
 (Br₂)₀ = 0.008240 M
 (HBr)₀ = 0.200 M
 (HClO₄) = 0.100 M
 50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.71		0.0
422	35.41	0.0307	9.8
1269	31.92	0.0296	24.7
2922	27.28	0.0290	44.4
4399	24.36	0.0295	56.9
6014	21.82	0.0311	67.7
7444	21.00	0.0285	71.2

Average rate constant (k_{app}) = 0.0297 ± 2.7%

Table 55

Run 62, DBPHBA_b in 80.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.004361 M

(Na₂S₂O₃) = 0.02185 N

(Br₂)₀ = 0.008168 M

(HBr)₀ = 0.100 M

(HClO₄) = 0.200 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.38		0.0
108.5	35.57	0.110	9.1
260.5	33.40	0.111	19.9
396.0	31.85	0.109	27.7
672	29.30	0.108	40.5
1241	25.75	0.106	58.3
2906	21.20	0.099	81.1

Average rate constant (k_{app}) = 0.107 ± 2.9%

Table 56

Run 61, DBPHBA_p in 80.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)_o = 0.004470 M

(Na₂S₂O₃) = 0.02185 N

(Br₂)_o = 0.008128 M

(HBr)_o = 0.0500 M

(HClO₄) = 0.250 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.20		0.0
11.0	36.30	(0.512)	4.4
40.0	34.58	0.436	12.8
87.5	32.21	0.423	24.4
139.0	30.25	0.410	34.0
286.0	26.50	0.384	52.3
476.5	23.60	0.366	66.5

Rate constant extrapolated to zero time (k_{app}) = 0.456

Table 57

Run 60, DBPHBA_p in 80.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.004163 M

(Na₂S₂O₃) = 0.02195 N

(Br₂)₀ = 0.008187 M

(HClO₄) = 0.300 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.30		0.0
9.8	26.35	13.04	57.7
23.2	25.27	11.32	79.2
47.2	21.90	9.83	91.7
88.5	18.70		98.1
2612	18.21		100.6

Rate constant extrapolated to zero time (k_{app}) \approx 15

are given of a series of runs in which the hydrogen ion concentration was held constant at 0.300 M and the bromide ion concentration was varied between zero and 0.300 M.

To demonstrate that TBP is indeed one of the products of the reaction, a solution 0.0066 M in DBPHBA and 0.0066 M in bromine in the solvent of Run 56, Table 48, was allowed to react for six days at 40°. From this mixture, 0.05 g. of TBP (55 per cent yield) was isolated. The TBP was identified by the melting point of the crude product, 88-90°, and by a mixed melting point with the product isolated from a DBSA reaction.

Finally, in Tables 49, 58 and 59, the results are recorded of the runs in which the effect of water concentration of the solvent was demonstrated. It is seen from these runs that a 10 per cent increase in the concentration of water will increase the rate constant 50 fold.

Phenols.--In order to compare the rate of bromination of DBSA with its parent phenol, 2,4-dibromophenol was brominated in 75.2 per cent acetic acid, 0.100 M in bromide ion and 0.100 M in hydrogen ion. The results of this run are shown in Table 60. This phenol was also brominated in 80.2 per cent acetic acid in Run 40 as shown in Table 61. When the methyl ether of 2,4-dibromophenol was subjected to the same conditions of bromination as those of Run 40, an initial reaction indicating bromination of about 2 per

Table 58

Run 71, DBPHBA_b in 75.0 per cent Acetic Acid-(b), 20.0°
 (DBPHBA)₀ = 0.004440 M (Na₂S₂O₃) = 0.02136 N
 (Br₂)₀ = 0.008042 M
 (HBr)₀ = 0.300 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.65		0.0
79.5	36.12	0.123	7.4
198.0	34.20	0.120	16.6
379.0	31.84	0.117	28.0
547.5	30.05	0.117	36.6
893.0	27.12	0.118	50.7
1266	25.07	0.115	60.5

Average rate constant (k_{app}) = 0.118 \pm 1.8%

Table 59

Run 72, DBPHBA_p in 70.0 per cent Acetic Acid-(b), 20.0°

(DBPHBA)₀ = 0.004441 M

(Na₂S₂O₃) = 0.02136 N

(Br₂)₀ = 0.007508 M

(HBr)₀ = 0.300 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	35.15		0.0
12.1	33.52	(0.914)	7.8
22.0	32.97	0.696	10.5
40.9	31.33	0.701	18.4
75.4	28.88	0.703	30.2
154.8	25.01	0.692	48.8
218.1	22.90	0.689	58.9

Average rate constant (k_{app}) = 0.696 ± 0.7%

Table 60

Run 33, 2,4-DBP in 75.2 per cent Acetic Acid-(b), 20.0°
 (DBP)₀ = 0.004878 M (Na₂S₂O₃) = 0.0202 N
 (Br₂)₀ = 0.003770 M
 (LiBr) = 0.100 M
 (HClO₄) = 0.100 M
 25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	10.30		0.0
6.8	4.95	30.25	51.9
13.6	3.11	30.00	69.8
29.7	1.56	28.22	84.9
55.5	0.78	25.83	92.4

Table 61

Run 40, 2,4-DBP in 80.2 per cent Acetic Acid-(b), 20.0°
 $(\text{DBP})_0 = 0.005124 \text{ M}$ $(\text{Na}_2\text{S}_2\text{O}_3) = 0.0200 \text{ N}$
 $(\text{Br}_2)_0 = 0.007472 \text{ M}$
 $(\text{LiBr}) = 0.1037 \text{ M}$
 $(\text{HClO}_4) = 0.100 \text{ M}$
 25.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	18.68		0.0
8.5	12.52	12.80	48.1
10.8	11.67	12.70	54.7
19.8	9.62	12.15	70.7
34.2	8.02	11.70	83.2
62.0	6.92	10.37	91.8
730.0	6.27		96.9
160.0 hrs.	6.22		97.3

cent of the ether was observed and this was followed by a virtually constant titer. The same results were observed in 50 per cent acetic acid having the same hydrogen and bromide ion concentrations. The results of these runs are shown in Table 62.

A series of runs similar to those made on DBSA and DBPHBA, in which the effect of hydrogen ion concentration and bromide ion concentration was determined, was also made on 2,6-DBP (the parent phenol of DBPHBA). Tables 63-65 show the results of holding the hydrogen ion concentration constant at 0.300 M and varying the bromide ion concentration between 0.100 M and 0.300 M. Tables 63 and 66-68 show the results of holding the bromide ion constant at 0.300 M and varying the concentration of hydrogen ion between 0.100 M and 0.300 M.

Table 62

Run 41, Methyl-(2,4-DBP) ether in 80.2 per cent
Acetic Acid-(b), 20.0°

Run 42, Methyl-(2,4-DBP) ether in 50.2 per cent
Acetic Acid-(b), 20.0°

(Ether)₀ = 0.005400 M (Run 41)

(Na₂S₂O₃) = 0.0200 N

(Ether)₀ = 0.005026 M (Run 42)

(LiBr) = 0.1037 M

(HClO₄) = 0.100 M

25.0 ml. of reaction mixture per flask

Run 41		Run 42	
Time elapsed (min.)	Titer (ml.)	Time elapsed (min.)	Titer (ml.)
(blank)	19.76	(blank)	19.28
8	19.68	14	19.05
26	19.62	32	19.05
97	19.58	121	18.90
220	19.49	256	19.00
681	19.50	1460	19.05
1410	19.58		

Table 63

Run 65, 2,6-DBP in 80.0 per cent Acetic Acid-(b), 20.0°

(DBP)₀ = 0.002605 M

(Na₂S₂O₃) = 0.02131 N

(Br₂)₀ = 0.004060 M

(HBr)₀ = 0.300 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	19.05		0.0
4.85	17.10	9.36	16.0
8.95	15.68	9.81	27.6
15.0	14.18	9.77	39.8
25.2	12.40	9.72	54.4
46.8	10.33	9.37	71.4
120.0	7.01		98.5

Average rate constant (k_{app}) = 9.61 ± 2.0%

Table 64

Run 66, 2,6-DBP in 80.0 per cent Acetic Acid-(b), 20.0°

(DBP)₀ = 0.002617 M

(Na₂S₂O₃) = 0.02131 N

(Br₂)₀ = 0.003753 M

(HBr)₀ = 0.200 M

(HClO₄) = 0.100 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	17.61		0.0
3.00	15.92	(13.76)	13.8
5.00*	14.83	-	-
5.00	14.90	14.48	22.1
10.07	12.99	14.66	37.6
14.03	11.88	14.75	46.7
18.57	10.90	14.74	54.6

Average rate constant (k_{app}) = 14.66 ± 0.6%

*This sample was allowed to stand 10 minutes before the iodine was titrated (ca. 30 seconds to 1 minute is the normal delay). The small drop in titer probably indicates reaction of iodine with the phenol, however, this reaction is obviously too slow to influence significantly the accuracy of the titrations.

Table 65

Run 67, 2,6-DBP in 80.0 per cent Acetic Acid-(b), 20.0°

(DBP)₀ = 0.002544 M

(Na₂S₂O₃) = 0.02131 N

(Br₂)₀ = 0.003714 M

(HBr)₀ = 0.100 M

(HClO₄) = 0.200 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	17.43		0.0
3.15	14.12	30.93	27.7
5.52	12.48	31.23	41.5
8.17	11.22	30.72	52.0
11.23	10.08	31.08	61.6
13.75	9.43	30.73	67.0
17.42	8.72	30.20	73.0

Average rate constant (k_{app}) = 30.82 ± 0.8%

Table 66

Run 68, 2,6-DBP in 80.0 per cent Acetic Acid-(b), 20.0°

(DBP)₀ = 0.002577 M

(Na₂S₂O₃) = 0.02131 N

(Br₂)₀ = 0.003729 M

(HBr)₀ = 0.200 M

(LiBr) = 0.100 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	17.50		0.0
4.90	15.22	12.30	18.9
10.00	13.57	12.03	32.5
15.00	12.38	11.82	42.3
20.03	11.39	11.90	50.5
30.00	10.04	11.67	61.7
40.00	9.04	11.77	70.0

Average rate constant (k_{app}) = 11.91 \pm 1.4%

Table 67

Run 70, 2,6-DBP in 80.0 per cent Acetic Acid-(b), 20.0°

(DBP)₀ = 0.002557 M

(Na₂S₂O₃) = 0.02136 N

(Br₂)₀ = 0.007982 M

(HBr)₀ = 0.150 M

(LiBr)₀ = 0.150 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	37.37		0.0
3.00	33.47	(17.45)	32.6
5.00	32.57	13.81	40.1
8.00	30.75	14.06	55.3
11.00	29.41	14.30	66.5
14.00	28.61	13.80	73.2
18.10	27.70	13.74	80.8

Average rate constant (k_{app}) = 13.94 ± 1.4%

Table 68

Run 69, 2,6-DBP in 80.0 per cent Acetic Acid-(b), 20.0°

(DBP)₀ = 0.002636 M

(Na₂S₂O₃) = 0.02131 N

(Br₂)₀ = 0.003687 M

(HBr)₀ = 0.100 M

(LiBr) = 0.200 M

50.0 ml. of reaction mixture per flask

Time elapsed (min.)	Titer (ml.)	Rate constant (l./mol.min.)	Per cent reaction
(blank)	17.30		0.0
3.0	15.08	19.17	17.9
5.0	14.03	18.56	26.4
8.0	12.80	17.97	36.4
11.0	11.80	17.78	44.5
14.0	11.00	17.63	50.9
18.0	10.12	17.56	58.0

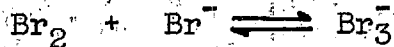
Average rate constant (k_{app}) = 18.11 ± 2.8%

CHAPTER VII

THE MECHANISM OF BROMINATION

3,5-Dibromo-4-hydroxybenzoic acid (DBPHBA).--The rate constants that have been tabulated in Chapter VI have all been calculated on the basis of a second order reaction, first order in bromine and first order in DBPHBA. This is justified on the grounds that at constant hydrogen ion and bromide ion concentrations the magnitude of the rate constant is comparatively independent of the variations in concentration of the DBPHBA and the bromine, e.g., see Tables 33, 34, and 40. That there are variations between runs is quite true, but it is these variations with hydrogen ion and bromide ion concentrations which are indicative of the true mechanism.

The variation in any one run is due to three factors: the first is that if there is an error in the initial concentration of either reagent, the integrated rate constants will deviate from constancy; the second is that due to the equilibrium



the bromide ion liberated in the reaction is continually altering the concentration of the brominating species;

and the third is that hydrogen ion is also being liberated in the reaction and the reaction is known to be retarded by increasing the hydrogen ion concentration.

Since the error in the initial concentrations is subject to only limited control (the technique and accuracy of the experimenter) and is not easily measurable, it must for the most part be ignored. It should be stated that the probable error of all concentrations except bromine is not greater than 0.5 per cent. The bromine concentration is determined by thiosulfate titration of the equivalent iodine and the end point is probably not known to within any better than 0.05 ml. and the total error in buret reading may amount to as much as 0.10 ml. Hence if a titer of 30 ml. is obtained, the error in bromine concentration would be roughly the same as that of the other compounds, however with smaller titers, the error may become somewhat larger.

Since it was shown very quickly (Table 32) that the rate is greatly slowed by added bromide ion, it is obvious that any rate constant calculated by using a stoichiometric bromine concentration and ignoring the bromide ion concentration is really only an apparent rate constant. When the bromide ion dependency was measured in the runs recorded in Tables 49 and 54-57, it was found that to a first approximation, the product of the apparent rate constant for a run

(the average of those obtained except in the cases where the initial bromide ion was less than 0.1 M, in which cases the constant was found by extrapolating to zero time) and the square of the stoichiometric initial bromide ion concentration is constant. This is shown in Table 69.

It was demonstrated that the rate of the reaction is dependent on the hydrogen ion concentration in Run 47, Table 42. It was shown quantitatively in Runs 54-58, Tables 49-53, that the product of the square of the initial hydrogen ion concentration times the apparent rate constant is a constant. The self ionization of acetic acid makes it rather difficult to know the exact hydrogen ion concentration when only a small amount of mineral acid has been added. By arbitrarily assuming that the ionization constant for 80 per cent acetic acid is half that of a dilute acetic acid solution, approximate corrections have been made on the initial hydrogen ion concentrations of the less acidic solvents. The hydrogen ion data are shown in Table 70.

The problem now becomes one of incorporating the bromide ion dependency and the hydrogen ion dependency into a rate equation and then considering various mechanisms which would give rise to the derived equation.

Since a dependency on the square of the bromide ion has been observed, it seems proper to assign part of this

Table 69

Effect of Bromide Ion Concentration on Reaction Rate
of DBPHBA

All runs were made in 80.0 per cent acetic acid,
0.300 M in hydrogen ion, at an ionic strength of 0.3 and at
20.0°.

k_{app} (l./mol.min.)	$(Br^-)_0$	$k_{app} \times (Br^-)_0^2$ (mol./l.min.)	Refer to Table
(15)	0.000		57
0.456	0.050	0.00114	56
0.111	0.100	0.00111	55
0.0297	0.200	0.00119	54
0.0135	0.300	0.00122	49

Table 70

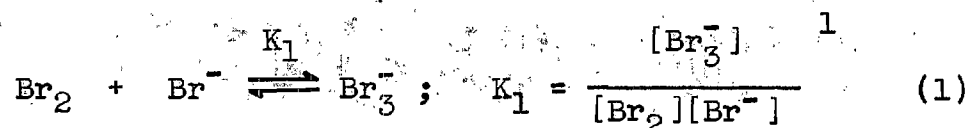
Effect of Hydrogen Ion Concentration on Reaction Rate
of DBPHBA

All runs were made in 80.0 per cent acetic acid
0.300 M in bromide ion (stoichiometric), at an ionic strength
of 0.3 and at 20.0°.

k_{app} (l./mol.min.)	$[H^+]_0$	$k_{app} \times [H^+]_0^2$ (mol./l.min.)	Refer to Table
0.0135	0.300	0.00122	49
0.0313	0.200	0.00125	50
0.120	0.100 + 0.001 ^a	0.00123	52
0.414	0.050 + 0.004	0.00121	53

^aApproximate hydrogen ion contribution from acetic
acid.

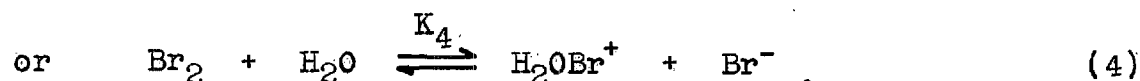
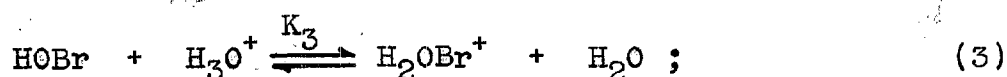
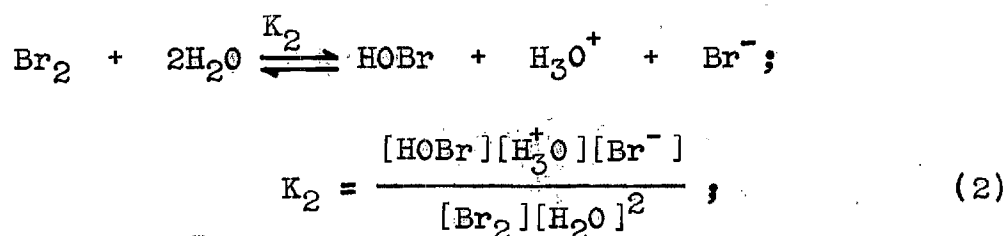
to the equilibrium of bromine and bromide with tribromide. Where it is desirable to distinguish between the stoichiometric concentration of a species and the actual concentration, the stoichiometric concentration will be designated by enclosure in parentheses, (), and the actual concentration by enclosure in square brackets []. Thus,



and hence:

$$[\text{Br}_2] = \frac{[\text{Br}_3^-]}{K_1 [\text{Br}^-]}.$$

If one considers the possible hydrolysis of bromine in an aqueous medium, a reason for the second power of bromide ion can be seen, i.e.,



¹The equilibrium constants defined in this chapter are not thermodynamic equilibrium constants but are constants expressed in concentrations.

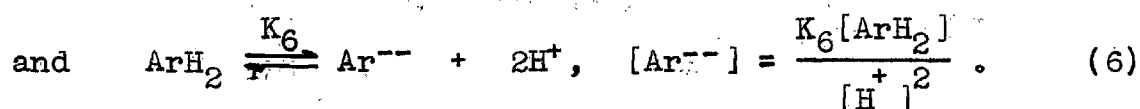
$$\text{Then } [\text{HOBr}] = \frac{K_2[\text{Br}_3^-][\text{H}_2\text{O}]^2}{K_1[\text{Br}^-]^2[\text{H}_3\text{O}^+]} \quad (3')$$

$$[\text{H}_2\text{OBr}^+] = \frac{K_3K_2[\text{Br}_3^-][\text{H}_2\text{O}]}{K_1[\text{Br}^-]^2} \quad (4')$$

$$\text{or } [\text{H}_2\text{OBr}^+] = \frac{K_4[\text{Br}_3^-][\text{H}_2\text{O}]}{K_1[\text{Br}^-]^2} \quad (4'')$$

Thus it is seen that the concentration of either HOBr or H_2OBr^+ is inversely proportional to the square of the bromide ion concentration. Since it is very difficult to distinguish H_2OBr^+ from Br^+ , the latter must also be included as a possibility. Because the expression for [HOBr] is inversely proportional to the hydrogen ion concentration, the hydrogen ion dependency of the reaction must be examined before a choice of mechanism can be made.

Table 70 shows a dependence of the rate on the square of the hydrogen ion concentration. Since DBPHBA is a dibasic acid, one or both of the acid hydrogen ions might be ionized in the reactive species, causing a first or second power hydrogen ion dependence respectively, i.e.,



Since both K_5 and K_6 are small, the actual concentration of the un-ionized acid, ArH_2 , can be replaced by the stoichiometric concentration.

Since the rate of the reaction, dx/dt , is independent of the equation used to calculate the rate constant, the following equations show how the apparent second order rate constant calculated on the basis of the stoichiometric bromine and stoichiometric DBPHBA concentrations can be related to actual rate constants which are dependent on the concentrations of either HOBr , H_2OBr^+ , or Br^+ , and an ionized species of DBPHBA.

$$\text{rate} = k_{\text{app}}(\text{ArH}_2)(\text{Br}_2) \quad (7)$$

$$\text{and if also} \quad \text{rate} = k_{\text{act}}[\text{H}_2\text{OBr}^+][\text{Ar}^{--}] \quad (8)$$

then from (6) and (4") above,

$$\text{rate} = \frac{k_{\text{act}}K_4[\text{Br}_3^-][\text{H}_2\text{O}][K_6(\text{ArH}_2)]}{K_1[\text{Br}^-]^2[\text{H}^+]^2} = \frac{k'_{\text{act}}[\text{Br}_3^-](\text{ArH}_2)}{K_1[\text{Br}^-]^2[\text{H}^+]^2}.$$

Then since

$$\text{rate} = k_{\text{app}}(\text{ArH}_2)(\text{Br}_2) = \frac{k'_{\text{act}}[\text{Br}_3^-](\text{ArH}_2)}{K_1[\text{Br}^-]^2[\text{H}^+]^2},$$

$$k'_{\text{act}} = \frac{K_1 k_{\text{app}}[\text{Br}^-]^2(\text{Br}_2)[\text{H}^+]^2}{[\text{Br}_3^-]} \quad (9)$$

In a similar way, this expression (equation 9) can be shown to hold for

$$\text{rate} = k_{\text{act}}[\text{HOBr}][\text{ArH}^-], \quad (10)$$

and since Br^+ is not readily distinguished from H_2OBr^+ , equation 9 holds equally for

$$\text{rate} = k_{\text{act}}[\text{Br}^+][\text{Ar}^{--}]. \quad (11)$$

The expression for k'_{act} can be evaluated if one knows K_1 (equation 1). This has been evaluated (see appendix) and the results of equation 9 are shown in Table 71.

Although distinction between bromination by HOBr and H_2OBr^+ cannot be made in this case, in other cases studied under similar conditions, the evidence is that H_2OBr^+ (or Br^+) is the reactive species. Bromination by HOBr can be distinguished from bromination by H_2OBr^+ (or Br^+) in those cases where ionization of the substrate is not involved in the hydrogen ion dependency of the rate (see equations 3' and 4"). It has been shown that in aqueous acidic solutions benzene (36), *o*-nitroanisole (36), sodium anisole *m*-sulfonate (34), sodium toluene-*u*-sulfonate and benzoic acid (80) are indeed brominated by H_2OBr^+ (or Br^+). The latter two were found not to react with vacuum distilled, bromide free hypobromous acid at a pH of 5.18.

80. D. N. Derbyshire and W. A. Waters, J. Chem. Soc., 1950, 564.

Table 71

Evaluation of k'_{act} from k_{app} for DBPHBA

$$k'_{act} = \frac{k_{app} K_1 [Br^-]^2 (Br_2) [H^+]^2}{[Br_3^-]}$$

 $K_1 = 93$ (in 80 per cent acetic acid)

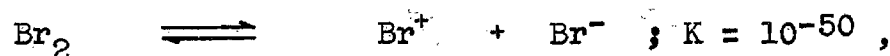
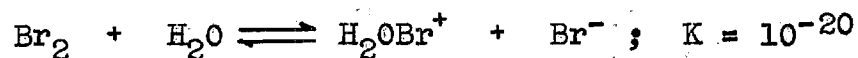
All runs were made in 80.0 per cent acetic acid, at an ionic strength of 0.3 and at 20.0°.

$(Br_2)_0$	$[Br^-]_0$	$[Br_3^-]_0$	$[H^+]_0$	k_{app}	k'_{act}	Refer. Table
0.00806	0.292	0.00777	0.300	0.0135	0.00101	70
0.00825	0.292	0.00795	0.200	0.0313	0.00103	70
0.00812	0.292	0.00783	0.101	0.120	0.00095	70
0.00816	0.292	0.00786	0.054	0.414	0.00101	70
0.00824	0.192	0.00780	0.300	0.0297	0.00097	69
0.00817	0.093	0.00732	0.300	0.111	0.00089	69
0.00813	0.043	0.00653	0.300	0.456	0.00089	69
0.00819	0.000	0.000	0.300	(15)	ca. 0.001 ^a	69

^aThis value was computed by assuming that the effective bromide ion concentration in the first 50 per cent of the reaction was constant and corresponded to the average stoichiometric bromide concentration of 0.001 M.

When the mixture was brought to pH 2 by perchloric acid, a reaction that was first order with respect to each reactant took place. It does not seem likely that a change in solvent from water to 80 per cent acetic acid would cause a significant change in the basicity of HOBr, and hence in the concentration of its conjugate acid, H_2OBr^+ . It is of course possible that some specific property of DBPHBA would favor bromination by HOBr, but this also seems rather unlikely.

Although it is difficult to distinguish between Br^+ and H_2OBr^+ , thermodynamic calculations by Bell and Gelles (81) of the equilibrium constant for the reactions in water at 25°,



suggest that the hydrated form of the bromonium ion is much more probable.

One disturbing factor is that in most of the other brominations that have been reported in which molecular bromine was used as the reagent, either in water, or acetic acid or in mixtures of the two, the kinetic evidence has always been that the brominating agent was Br_2 and the reaction may or may not have been complicated by the formation

81. R. P. Bell and E. Gelles, J. Chem. Soc., 1951, 2734.

of Br_3^- or $(\text{Br}_2)_2$ (31, 35, 36, 38, and 80). A consideration of a mechanism involving Br_2 which is also compatible with all of the kinetic data will be considered following the discussion of the kinetics of the reaction involving DBSA.

What can be definitely concluded from the present kinetic study is that for finite bromide ion concentrations the transition state complex for the slowest step of the reaction contains the elements of Br^+ and the di-ionized form of DBPHBA.

2,6-Dibromophenol (2,6-DBP).---The possibility was considered that the brominative decarboxylation reaction occurred by way of a preliminary fast displacement of the carboxyl group by a hydrogen ion to give a dibromophenol. If this were the case, the rate of bromination measured for the hydroxybenzoic acid would be the same as the rate of bromination of its corresponding dibromophenol. By comparing the data in Tables 63 and 49 and in Tables 60 and 18 it is seen that in both cases the phenol is brominated over 100 times faster than the corresponding hydroxybenzoic acid. This evidence, and the evidence recorded in Table 41 in which it is shown that the rate of evolution of carbon dioxide from DBPHBA is very nearly the same as its rate of bromine consumption, clearly disprove that the carboxyl group is displaced in a fast step by a hydrogen ion.

The possibility of the carboxyl group being displaced in a slow rate determining step by hydrogen ion is also disproved by the fact that the reactions all show a kinetic dependence on the concentration of bromine.

In order to compare the mechanism of the displacement of the carboxyl group by bromine with the displacement of a hydrogen ion by bromine, the hydrogen ion and bromide ion dependence of the rate of bromination of 2,6-DBP were determined. The information obtained is tabulated in Tables 72 and 73.

Table 72 shows that the hydrogen ion dependence is somewhere in between zero and first power. Since the hydrogen ion dependence of a phenol would reasonably be expected to be related to its ionization to phenoxide ion, the interpretation of this data is that both phenol and phenoxide ion are being brominated.

Table 73 shows that to a reasonably close approximation, the rate is inversely dependent on the first power of the stoichiometric bromide ion concentration. This would then indicate that Br_2 is the brominating agent since the actual concentration of Br_2 is a function of the bromide ion concentration.

These data can then be expressed by the rate equation

$$\text{rate} = k[\text{Br}_2][(\text{PhOH}) + \frac{k'(\text{PhOH})}{[\text{H}^+]}], \quad (12)$$

Table 72

Effect of Hydrogen Ion Concentration on the Rate of
Bromination of 2,6-DBP

All runs were made in 80.0 per cent acetic acid,
0.300 M in bromide ion (stoichiometric), at an ionic strength
of 0.3 and at 20.0°.

k_{app} (l./mol.min.)	$[H^+]_0$	$k_{app} \times [H^+]_0$ (min. ⁻¹)	Refer to Table
9.61	0.300	2.883	63
11.91	0.200	2.382	66
13.94	0.150 + 0.001 ^a	2.104	67
18.11	0.100 + 0.001 ^a	1.829	68

^aApproximate correction for contribution of acetic
acid to hydrogen ion concentration.

Table 73

Effect of Bromide Ion Concentration on the Rate of
Bromination of 2,6-DBP

All runs were made in 80.0 per cent acetic acid,
0.300 M in hydrogen ion, at an ionic strength of 0.3 and at
20.0°.

k_{app} (l./mol.min.)	$(Br^-)_0$	$k_{app} \times (Br^-)_0$ (min. ⁻¹)	$k_{app} \times (Br^-)_0^2$ (mol./l.min.)	Ref. Table
9.61	0.300	2.88	0.864	63
14.7	0.200	2.94	0.588	64
30.8	0.100	3.08	0.308	65

and equating this to equation 7 and incorporating equation 1,

$$k_{app}(\text{PhOH})(\text{Br}_2) = k \frac{[\text{Br}_3^-]}{K_1[\text{Br}^-]} \left[(\text{PhOH}) + \frac{k'(\text{PhOH})}{[\text{H}^+]} \right].$$

$$\text{Then } \frac{k_{app}K_1(\text{Br}_2)[\text{Br}^-]}{[\text{Br}_3^-]} = k + \frac{kk'}{[\text{H}^+]} = F. \quad (13)$$

By plotting F vs. $1/[\text{H}^+]$ and evaluating by the method of least squares, it is possible to evaluate k and k' . Table 74 shows the values of F and the plot is shown in Fig. 2. From this plot it is found that $k' = 0.207 \text{ mol./l.}$ and $k = 169 \text{ l./mol.min.}$ The mean deviation of F is ± 9 . The distribution of points at the hydrogen ion concentration of 0.300 M is probably due to salt effects, since salt effects are expected to be somewhat specific in 80 per cent acetic acid at high ionic strength.

It is of interest to compare these observations on the bromination of phenol with those of Robertson, de la Mare, and Swedlund (82). They found that in glacial acetic acid at 25° , a solution 0.00100 M in phenol and 0.00100 M in bromine showed no alteration in rate of reaction upon making the solution 0.0074 M in sulfuric acid. They also report that phenol is brominated 100 times as fast as anisole. Both of these observations appear contrary to what has been

82. P. W. Robertson, P. B. de la Mare, and B. E. Swedlund, J. Chem. Soc., 1953, 782.

Table 74

Evaluation of F

$$F = k + \frac{kk'}{[H^+]} = \frac{k_{app} K_1 (Br_2)_o [Br^-]_o}{[Br_3^-]_o}$$

$$K_1 = 93$$

All runs were made in 80.0 per cent acetic acid at an ionic strength of 0.3 and at 20.0°.

$[H^+]_o$	$(Br_2)_o$	$[Br_3^-]_o$	$[Br^-]_o$	$1/[H^+]_o$	F	Refer to Table
0.300	0.00406	0.00392	0.296	3.33	274	73
0.300	0.00375	0.00356	0.196	3.33	282	73
0.300	0.00371	0.00334	0.0967	3.33	306	73
0.200	0.00373	0.00360	0.296	5.00	340	72
0.151	0.00798	0.00770	0.292	6.63	392	72
0.101	0.00369	0.00356	0.296	9.90	516	72

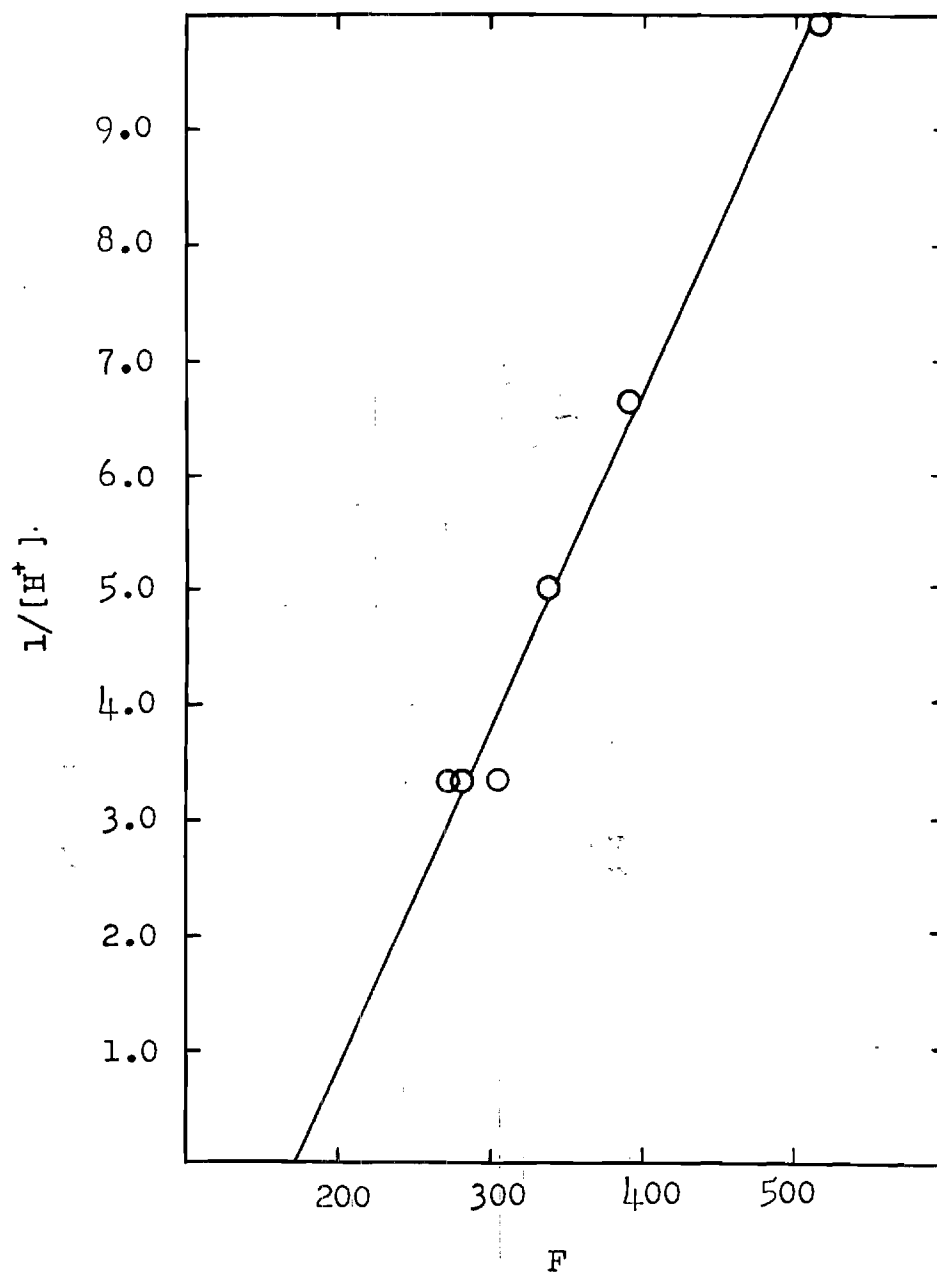


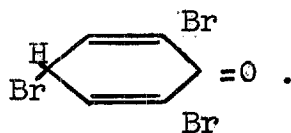
Fig. 2, 2,6-Dibromophenol

observed in this study. If this same ratio is used to compare methyl 2,4-dibromophenyl ether with 2,4-dibromophenol, the conclusion would be that the ether should have been brominated at the same rate as DBSA, however it was apparently not brominated at all.

Robertson claims that phenol is too weak an acid to be ionized measurably in acetic acid, and thus discounts the great reactivity of phenoxide ion as the reason for the difference in rate between anisole and phenol. He explains the difference on the basis of the hyperconjugation possible in phenol which is not possible in anisole, e.g.,

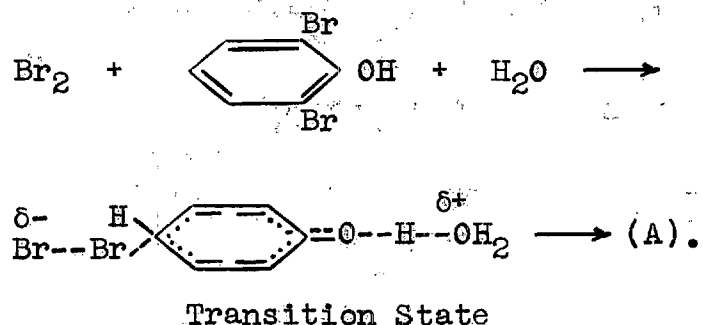


It is possible that by the inductive effect of the two bromine atoms substituted in DBP (either 2,4 or 2,6), that the acidity of the phenol is increased sufficiently for it to be measurably ionized, especially in aqueous acetic acid. If this is the case, the reactivity of the dibromophenol would be increased relative to that of the methyl dibromophenyl ether. The reactive intermediate in the bromination of 2,6-DBP is probably



(A)

Rather than a hyperconjugation effect, it seems more pleasing to suggest that the un-ionized phenol is brominated in a concerted reaction, e.g.,



The reason for the unreactivity of the phenol ether is not entirely clear, especially since many phenol ethers are readily brominated and have been studied (31, 38, 82). Although substituted halogens do increase the acidity of phenols by their inductive effect, and thus increase their reactivity, this same effect opposes the desired electromeric effect in the absence of an ionizable proton, resulting in a net deactivation. It is not felt that steric effects can be important, especially in the case of the ether of 2,4-DBP because in assuming the quinoidal transition state, the methyl group is not prohibited from lying in the plane of the ring.

3,5-Dibromo-2-hydroxybenzoic acid (DBSA).--The hydrogen ion dependence and the bromide ion dependence of the rate of bromination of DBSA are shown respectively in Tables 75 and 76. Empirically it is seen that the dependence is somewhere

between the first and second powers of the concentrations of both ions. It was hoped that DBSA and DBPHBA would both show the same dependence and could thus both be assumed to react with bromine by the same mechanism. It was not surprising that the hydrogen ion dependence might be different due to the possibility of hydrogen bonding between the hydroxyl group and the carbonyl group, but the different bromide ion dependence was unexpected.

The fact that the dependence of bromide ion is a non-integral power suggests that the bromide ion is involved in a reversible step. The following mechanism has been postulated to account for this:

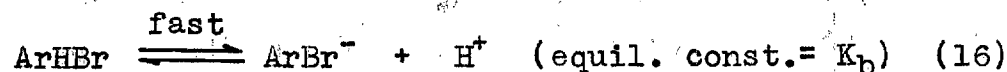
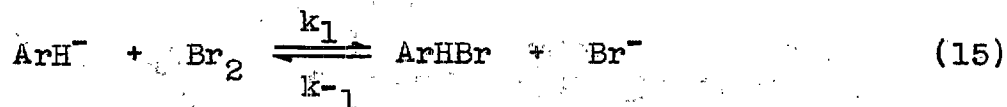
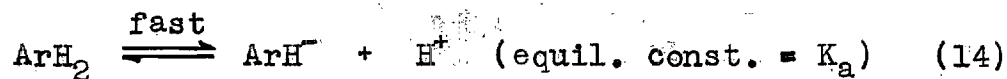


Table 75

Effect of Hydrogen Ion Concentration on the Rate of
Bromination of DBSA

All runs were made in 75.0 per cent acetic acid
0.300 M in bromide ion (stoichiometric) at an ionic strength
of 0.3 and at 20.0°.

k_{app} (l./mol.min.)	$[H^+]_0$	$k_{app} \times [H^+]_0$ (min. ⁻¹)	$k_{app} \times [H^+]_0^2$ (mol./l.min.)	Refer to Table
0.0142	0.300	0.00426	0.001278	22
0.0257	0.200	0.00514	0.001028	26
0.0689	0.101 ^a	0.00696	0.000704	27

^aAcetic acid is assumed to contribute about 0.001 M
to the total hydrogen ion concentration.

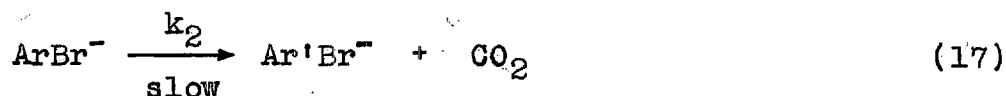
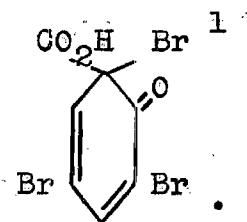
Table 76

Effect of Bromide Ion Concentration on the Rate of
Bromination of DBSA

All runs were made in 75.0 per cent acetic acid,
0.300 M in hydrogen ion, at an ionic strength of 0.3 and
at 20.0°.

k_{app} (l./mol.min.)	$(Br^-)_0$	$k_{app} \times (Br^-)_0$ (min. ⁻¹)	$k_{app} \times (Br^-)_0^2$ (mol./l.min.)	Ref. Table
0.0142	0.300	0.00426	0.00128	22
0.0264	0.200	0.00528	0.00106	21
0.0825	0.100	0.00825	0.000825	20
0.215	0.050	0.0108	0.000538	19
0.830	0.0100	0.00830	0.000083	24
1.65	0.000			25

where ArHBr is the reactive intermediate with the structure



Assume that ArHBr is formed in low concentration relative to ArH₂, then on the basis of the steady state approximation

$$k_1[\text{ArH}^-][\text{Br}_2] = k_1' \frac{(\text{ArH}_2)}{[\text{H}^+]} [\text{Br}_2] ; \quad (k_1' = k_1 K_a)$$

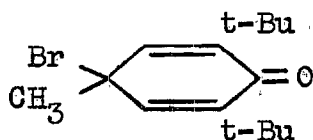
$$= k_{-1}[\text{Br}^-][\text{ArHBr}] + k_2[\text{ArBr}^-]$$

$$= k_{-1}[\text{Br}^-][\text{ArHBr}] + k_2' \frac{[\text{ArHBr}]}{[\text{H}^+]} ;$$

$$(k_2' = k_2 K_b)$$

$$[\text{ArHBr}] = \left(k_1' [\text{Br}_2] \frac{(\text{ArH}_2)}{[\text{H}^+]} \right) / \left(k_{-1} [\text{Br}^-] + \frac{k_2'}{[\text{H}^+]} \right) ;$$

¹It is of interest to compare this intermediate with the compound stated to have been isolated by G. M. Coppinger and T. W. Campbell, J. Am. Chem. Soc., 75, 734 (1953).



$$\text{rate} = \frac{k_2' k_1 [\text{Br}_2] (\text{ArH}_2)}{[\text{H}^+]^2 (k_{-1} [\text{Br}^-] + k_2' / [\text{H}^+])} = k_{\text{app}} (\text{Br}_2) (\text{ArH}_2)$$

$$\frac{k_{\text{app}} (\text{Br}_2)}{[\text{Br}_2]} = \frac{k_2' k_1}{[\text{H}^+]^2 (k_{-1} [\text{Br}^-] + k_2' / [\text{H}^+])}$$

Substituting for $[\text{Br}_2]$ from equation 2 and rearranging,

$$\frac{k_{\text{app}} K_1 (\text{Br}_2) [\text{Br}^-] [\text{H}^+]}{[\text{Br}_3^-]} = \frac{k_1}{\frac{k_{-1}}{k_2'} [\text{Br}^-] [\text{H}^+] + 1} = \Psi \quad (18)$$

$$\frac{1}{\Psi} = \frac{k_{-1}}{k_2' k_1} [\text{Br}^-] [\text{H}^+] + \frac{1}{k_1'} \quad (19)$$

A tabulation of the values of Ψ and the related hydrogen and bromide ion concentrations is shown in Table 77. A plot of $1/\Psi$ vs. $[\text{Br}^-][\text{H}^+]$ was made and then evaluated by the method of least squares and is shown in Fig. 3. The slope of the line is $k_{-1}/k_2' k_1' = 82.2 \text{ l.}^2 \text{ min.} / \text{mol}^2$. The intercept at $[\text{Br}^-][\text{H}^+]$ equal zero is $1/k_1'$ and from this $k_1' = 0.42 \text{ min.}^{-1}$. From this then is obtained $k_{-1}/k_2' = 34.6 \text{ l.}^2 / \text{mol.}^2$ which is an index of the reversibility of the reaction. The mean deviation of $1/\Psi$ is ± 0.27 .

It might be suggested that perhaps a mixed mechanism of the type

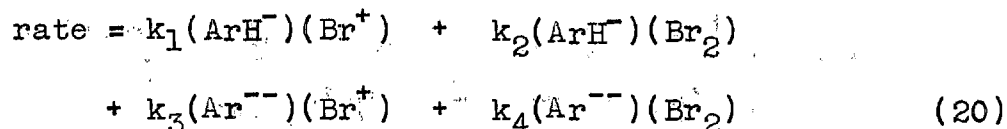


Table 77

Evaluation of Ψ for DBSA

$$\Psi = \frac{k_{app} K_1 (Br_2)_o [Br^-]_o [H^+]_o}{[Br_3^-]}$$

$$K_1 = 85$$

All runs were made in 75.0 per cent acetic acid at an ionic strength of 0.3 and at 20.0°.

k_{app}	$[H^+]_o$	$[Br^-]_o$	$[Br_3^-]_o$	$(Br_2)_o$	Ψ	Refer to Table
0.0142	0.300	0.292	0.008368	0.008368	0.1098	76
0.0264	0.300	0.193	0.007488	0.007954	0.138	76
0.0825	0.300	0.093	0.007075	0.007973	0.2205	76
0.215	0.300	0.044	0.006300	0.007996	0.306	76
0.830	0.300	0.0070	0.002968	0.007932	0.396	76
1.65	0.300	0.000	0.000	0.007942	0.495 ^a	76
0.0257	0.200	0.292	0.007656	0.007964	0.1327	75
0.0689	0.101	0.292	0.007593	0.007898	0.1801	75

^aAt zero bromide ion concentration, Ψ is equal to k_{app} times the hydrogen ion concentration.

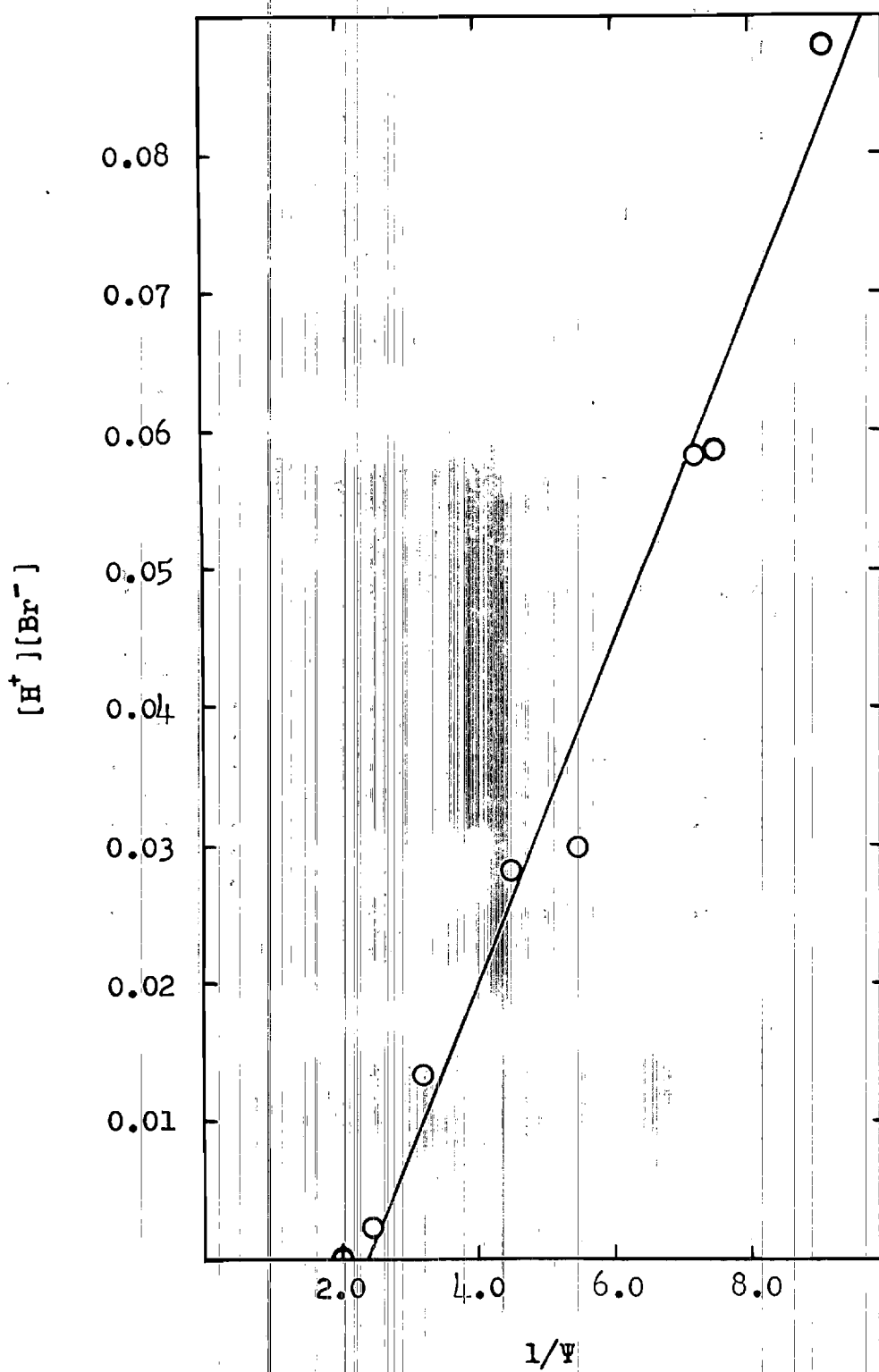


Fig. 3, Dibromosalicylic Acid

could qualitatively account for the hydrogen and bromide ion dependence. At constant hydrogen ion concentration this would resolve into

$$\begin{aligned} \text{rate} &= k_1(\text{ArH}_2) \frac{[\text{Br}_3^-]}{[\text{Br}^-]^2} + k_2'(\text{ArH}_2) \frac{[\text{Br}_3^-]}{[\text{Br}^-]} \\ &= \frac{(\text{ArH}_2)[\text{Br}_3^-]}{[\text{Br}^-]} \left[\frac{k_1}{[\text{Br}^-]} + k_2' \right]. \end{aligned}$$

And since $\text{rate} = k_{\text{app}}(\text{ArH}_2)(\text{Br}_2)$,

$$\begin{aligned} k_{\text{app}}(\text{ArH}_2)(\text{Br}_2) &= \frac{(\text{ArH}_2)[\text{Br}_3^-]}{[\text{Br}^-]} \left[\frac{k_1}{[\text{Br}^-]} + k_2' \right] \\ \frac{k_{\text{app}}(\text{Br}_2)[\text{Br}^-]}{[\text{Br}_3^-]} &= \frac{k_1}{[\text{Br}^-]} + k_2' = G. \end{aligned} \quad (21)$$

If G is plotted against $1/[\text{Br}^-]$, Fig. 4 is obtained. The data to plot Fig. 4 is shown in Table 78. It is obvious from the plot that the observed data do not fit this mechanism.

While it is felt that the mechanism resulting in equation 18 is very probably the correct treatment for the course of bromine and bromide ion, it can be shown that another course is equally satisfactory for the hydrogen ion. Thus if the reaction were initiated on the un-ionized hydroxybenzoic acid:

Table 78

Evaluation of G for DBSA

$$G = \frac{k_{app}(Br_2)_o [Br^-]_o}{[Br_3^-]_o}$$

All runs were made in 75.0 per cent acetic acid, 0.300 M in hydrogen ion, at an ionic strength of 0.3 and at 20.0°. Refer to Table 76.

k_{app}	$[Br^-]_o$	$[Br_3^-]_o$	$(Br_2)_o$	G
0.0142	0.292	0.008044	0.008368	0.00431
0.0264	0.193	0.007488	0.007954	0.00541
0.0825	0.093	0.007075	0.007973	0.00865
0.215	0.044	0.006300	0.007996	0.0120
0.830	0.0070	0.002968	0.007932	0.0155
1.65	0.000	0.0000	0.007942	0.0194 ^a

^aAt zero bromide ion concentration, G reduces to k_{app}/K_1 since $(Br_2)_o$ then is the same as $[Br_2]_o$. $K_1 = 85$.

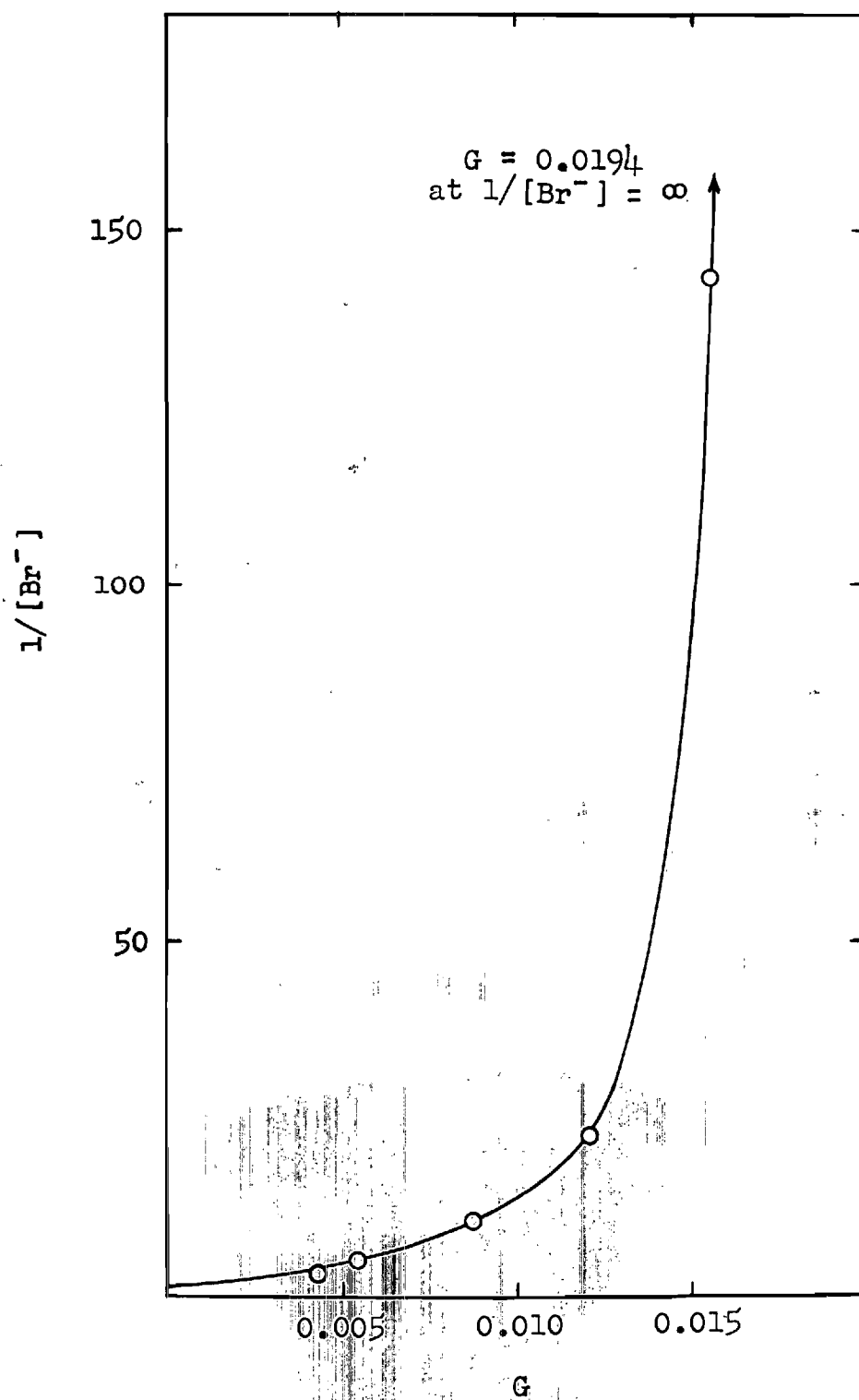
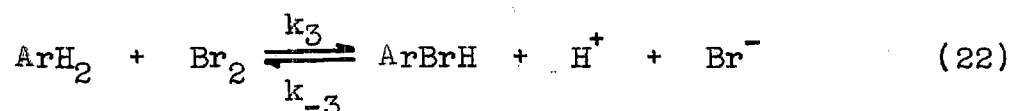
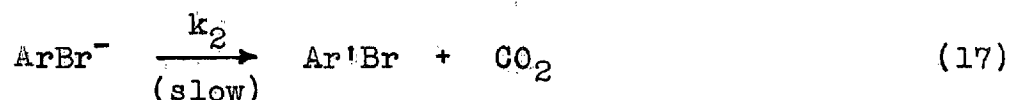
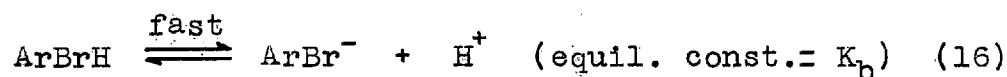


Fig. 4, Dibromosalicylic Acid



and then as previously discussed,



Then

$$\begin{aligned} k_3(\text{ArH}_2)[\text{Br}_2] &= k_{-3}[\text{ArBrH}][\text{H}^+][\text{Br}^-] + k_2[\text{ArBr}^-] \\ &= k_{-3}[\text{ArBrH}][\text{H}^+][\text{Br}^-] + k_2' \frac{[\text{ArBrH}]}{[\text{H}^+]}; \quad k_2' = k_2 K_b \end{aligned}$$

$$[\text{ArBrH}] = \frac{k_3(\text{ArH}_2)[\text{Br}_2]}{k_{-3}[\text{H}^+][\text{Br}^-] + \frac{k_2'}{[\text{H}^+]}}$$

$$\text{rate} = k_{\text{app}}(\text{ArH}_2)(\text{Br}_2) = \frac{k_2' k_3(\text{ArH}_2)[\text{Br}_2]}{k_{-3}[\text{H}^+]^2[\text{Br}^-] + k_2'}$$

$$\frac{k_{\text{app}} K_1 (\text{Br}_2) [\text{Br}^-]}{[\text{Br}_2]} = \frac{1}{\frac{k_{-3}}{k_2' k_3} [\text{H}^+]^2 [\text{Br}^-] + \frac{1}{k_3}} = Q \quad (23)$$

If $1/Q$ is plotted against $[\text{H}^+]^2[\text{Br}^-]$ and evaluated by the method of least squares, a straight line is obtained as shown in Fig. 5. The information required to make the plot

Table 79

Evaluation of Q for DBSA

$$Q = \frac{k_{\text{app}} K_1 (\text{Br}_2) [\text{Br}^-]}{[\text{Br}_3^-]} = \Psi / [\text{H}^+]$$

$$K_1 = 85$$

All runs were made in 75.0 per cent acetic acid at an ionic strength of 0.3 and at 20.0°. See Table 77.

$[\text{H}^+]_0$	$[\text{Br}^-]_0$	$[\text{H}^+]_0^2 [\text{Br}^-]_0$	Ψ	Q
0.300	0.292	0.0263	0.1098	0.366
0.300	0.193	0.0174	0.138	0.460
0.300	0.093	0.00837	0.2205	0.735
0.300	0.044	0.00396	0.306	1.020
0.300	0.0070	0.00063	0.396	1.321
0.300	0.000	0.000	0.495	1.650
0.200	0.292	0.0117	0.1327	0.663
0.101	0.292	0.00298	0.1801	1.785

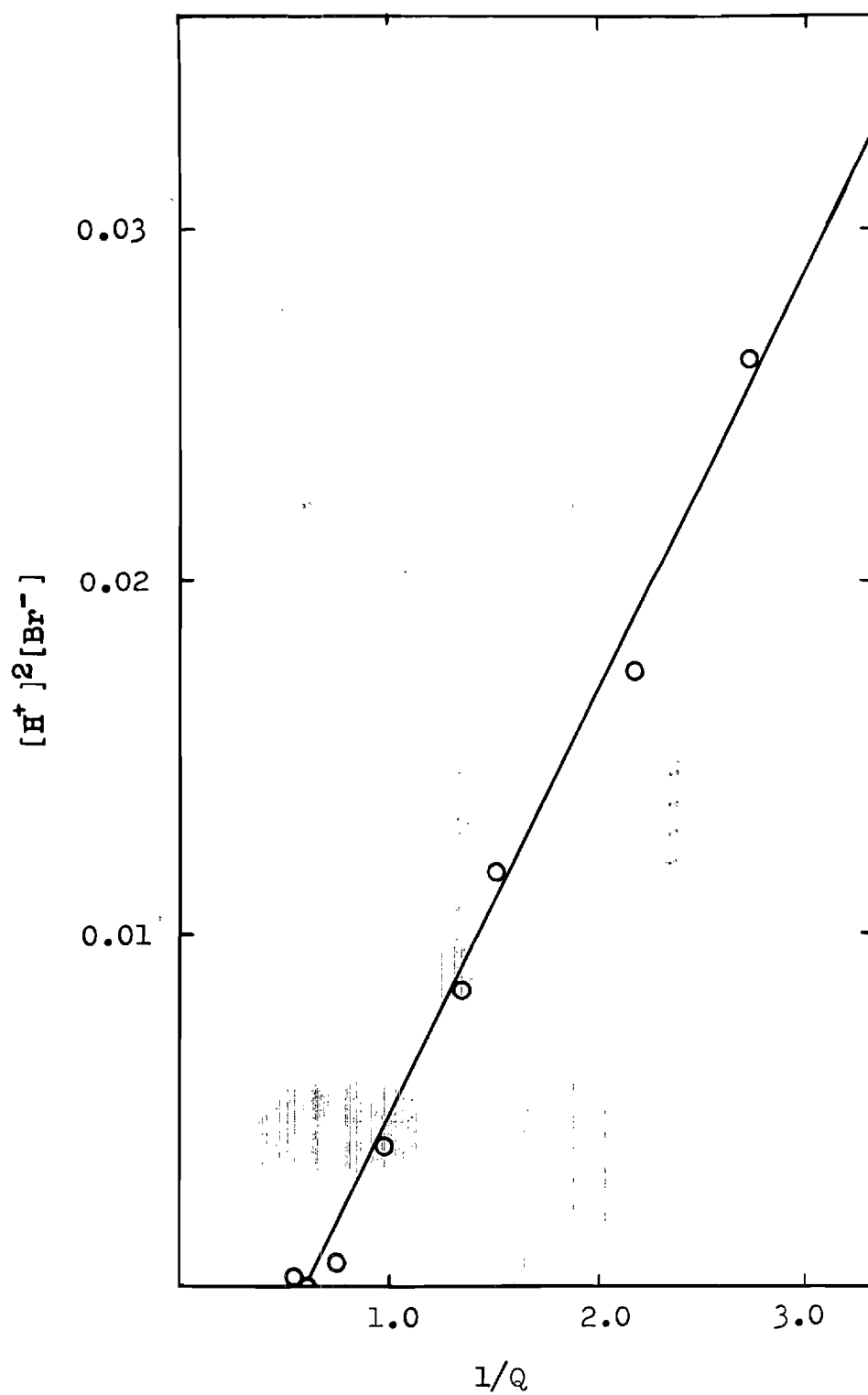


Fig. 5, Dibromosalicylic Acid

is given in Table 79. From this plot it is found that $k_3 = 1.7 \text{ l./mol.min.}$ and the reversibility ratio k_{-3}/k_2' is $143 \text{ l.}^3/\text{mol.}^3$. The mean deviation of $1/Q$ is ± 0.11 .

Since it has been shown that mechanisms involving an initial bromination of either the un-ionized or mono-ionized DBSA will satisfy the kinetic data that has been obtained, it is immediately suggested that analogously to the bromination of phenols, both species are reacting concurrently to give the common intermediate. If equations 15 and 22 are considered then to both give the common species ArBrH which decomposes in the irreversible step represented by equation 17, the kinetic expression resulting from such a consideration is

$$\frac{k_{\text{app}} K_1 (\text{Br}_2) [\text{Br}^-] [\text{H}^+]}{[\text{Br}_3^-]} = \frac{k_1' + k_3 [\text{H}^+]}{\frac{k_{-1}}{k_2'} [\text{H}^+] [\text{Br}^-] + \frac{k_{-3}}{k_2'} [\text{H}^+]^2 [\text{Br}^-] + 1} \quad (24)$$

Unfortunately, not enough experimental data is available to test this mechanism; however, it can be seen that at zero bromide ion concentration, this expression would reduce to the same form as that shown by equation 13 which represents the simultaneous bromination of phenol and phenoxide ion.

In order to establish the exact course of loss of

hydrogen ions in the reaction, the kinetics must be studied at zero bromide ion concentration. It can be seen from equations 15 and 22 for the steps involving combination with bromine, that in the absence of bromide ion, these steps would be irreversible and hence rate controlling. Then by measuring the rate of the reaction at varying hydrogen ion concentrations, the kinetic order with respect to hydrogen ion could be determined.

Rediscussion of DBPHBA.--The mechanism for DBSA represented by Fig. 3, while kinetically more complex than that proposed for DBPHBA, seems much more probable in the light of our knowledge of brominations and bromine species. Since it would seem reasonable that the same mechanism might apply to both compounds, a similar treatment was made with the data obtained for DBPHBA. Table 80 is a tabulation of Ψ and the related hydrogen and bromide ion concentrations for DBPHBA. A plot of $1/\Psi$ vs. $[\text{Br}^-][\text{H}^+]$ was made and evaluated by the method of least squares and is shown in Fig. 6. From the slope of this curve it is found that $k_{-1}/k_2'k_1'$ equals $97 \text{ l.}^2\text{min.}/\text{mol.}^2$. From the intercept at $[\text{Br}^-][\text{H}^+]$ equal to zero, it is determined that $k_1' = 6.3 \text{ min.}^{-1}$, and from these two that the reversibility ratio, $k_{-1}/k_2' = 612 \text{ l.}^2/\text{mol.}^2$. The mean deviation of $1/\Psi$ is ± 0.13 . Thus it is seen that although the concept of bromination by Br^+ or H_2OBr^+ adequately fits the data, the concept of an initial

Table 80

Evaluation of Ψ for DBPHBA

$$\Psi = \frac{k_{app} K_1 (Br_2)_o [Br^-]_o [H^+]_o}{[Br_3^-]_o}$$

$$K_1 = 93$$

All runs were made in 80.0 per cent acetic acid at an ionic strength of 0.3 and at 20.0°.

k_{app}	$[H^+]_o$	$[Br^-]_o$	$[Br_3^-]_o$	$(Br_2)_o$	Ψ	Refer to Table
0.0135	0.300	0.292	0.00777	0.00806	0.114	69
0.0297	0.300	0.192	0.00780	0.00824	0.168	69
0.111	0.300	0.093	0.00732	0.00817	0.321	69
0.456	0.300	0.043	0.00653	0.00813	0.682	69
(15)	0.300	0.000	0.00000	0.00819	4.5 ^a	69
0.0313	0.200	0.292	0.00795	0.00825	0.175	70
0.120	0.101	0.292	0.00783	0.00812	0.341	70
0.414	0.054	0.292	0.00786	0.00816	0.666	70

^aAt zero bromide ion concentration, Ψ is equal to k_{app} times the hydrogen ion concentration.

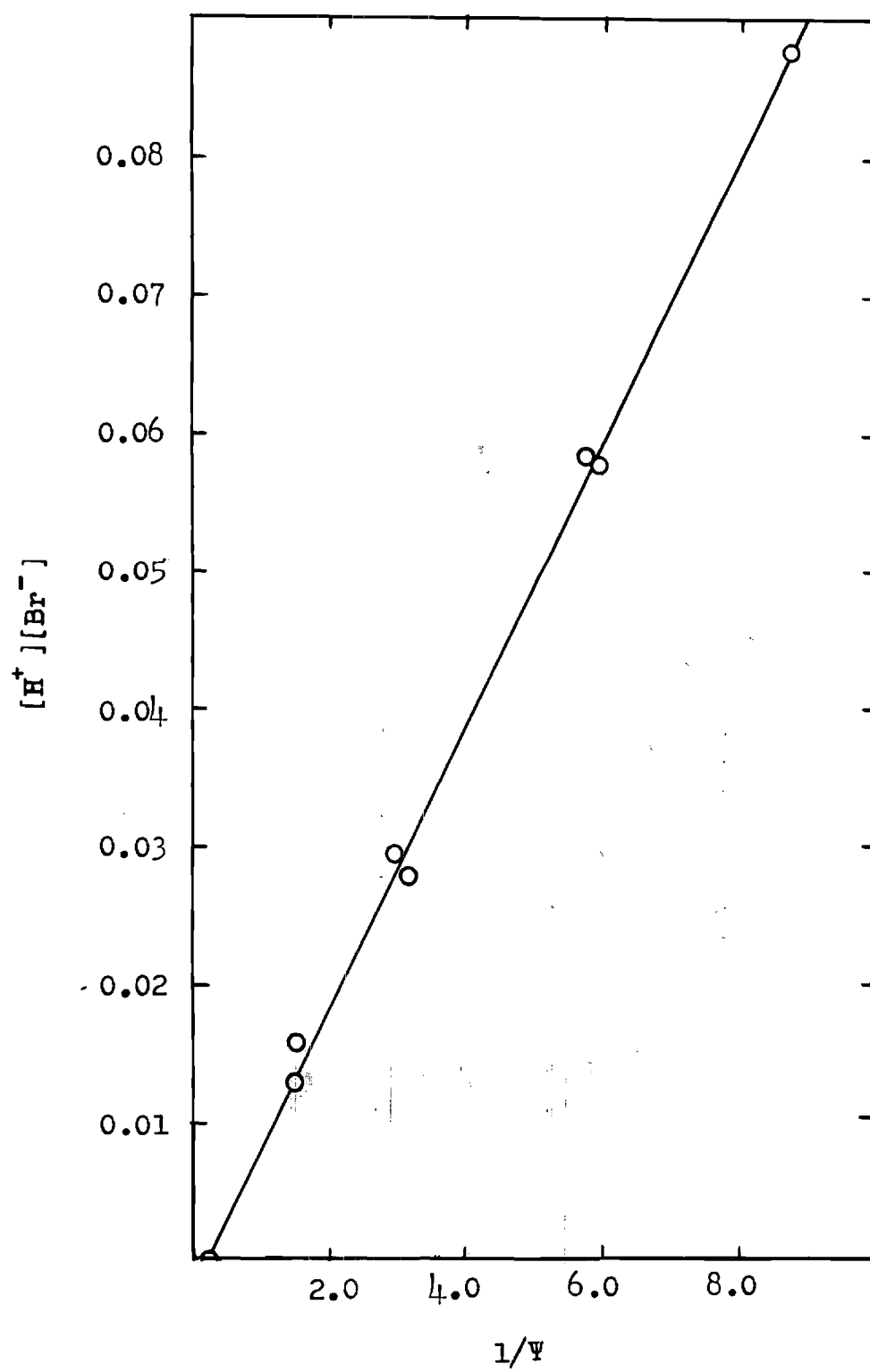


Fig. 6, Dibromo-p-hydroxybenzoic Acid

reversible step involving bromide ion also satisfactorily explains all of the data. Since the reversibility ratio, k_{-1}/k'_2 , is some 20 times greater for DBPHBA than for DBSA, it is understandable that empirically the bromide ion dependence for DBPHBA should seem more nearly second order. It can be seen mathematically from equation 18,

$$\Psi = \frac{k_{app} K_1 (Br_2) [Br^-] [H^+]}{[Br_3^-]} = \frac{k_1}{\frac{k_{-1}}{k'_2} [Br^-] [H^+] + 1}, \quad (18)$$

that when k_{-1}/k'_2 is much larger than 1,

$$\frac{k_{app} K_1 (Br_2) [Br^-]^2 [H^+]^2}{[Br_3^-]} = \text{constant}$$

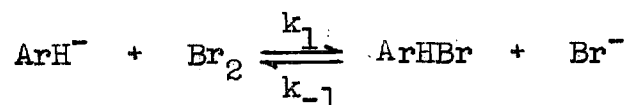
and corresponds to equation 9. For this relation to hold absolutely, the plot of $1/\Psi$ vs. $[H^+][Br^-]$, Fig. 6, should pass through the origin. Since the value of $1/\Psi$ at the intercept is 0.16 ± 0.13 , it is seen that almost within the experimental error, this may be the case. To the extent that the true intercept lies within the range of 0.16 ± 0.13 , the reversibility ratio is between 335 and 3233! A more accurate determination of the rate at low bromide ion concentrations is necessary before a distinction can be made between the two mechanisms for DBPHBA.

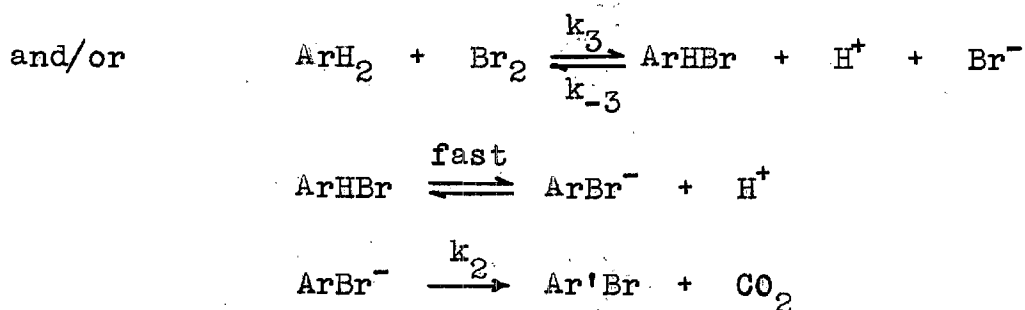
CHAPTER VIII

CONCLUSIONS

The rates of bromination of 2,4-dibromophenol, 2,6-dibromophenol, 3,5-dibromo-2-hydroxybenzoic acid, and 3,5-dibromo-4-hydroxybenzoic acid were measured in aqueous acetic acid and found at constant bromide and hydrogen ion concentration to be first order in bromine and first order in substrate. It was observed that the rate of bromination of 2,6-dibromophenol was inversely dependent on the first power of the bromide ion concentration and showed an inverse hydrogen ion dependence between zero and first power. This evidence is consistent with a mechanism in which bromine molecules react with both ionized and un-ionized phenol.

The dibromosalicylic acid was found to show an inverse hydrogen ion dependence and an inverse bromide ion dependence between first and second power. This data is found to be compatible with a mechanism involving a reversible attack by bromine on either the un-ionized or mono-ionized form of the acid, followed by an irreversible decarboxylation step, thus





The information available does not make possible a distinction between the first two equations nor does it give enough evidence to confirm a simultaneous bromination of the unionized and mono-ionized forms to give the same reactive intermediate. The kinetic data do not fit a mechanism involving simultaneous attack by Br^+ or H_2OBr^+ and Br_2 .

The rate of bromination of dibromo-p-hydroxybenzoic acid was found to be inversely dependent on the second power of both hydrogen ion and bromide ion concentrations. This is consistent with a mechanism of the type proposed above for dibromosalicylic acid. However it was found that the reversibility ratio, k_{-1}/k'_2 , is somewhere between 10 and 1000 times greater for the para compound than for the ortho compound. The kinetics is also consistent with a mechanism in which Br^+ or H_2OBr^+ reacts with the di-ionized form of the acid, thus



CHAPTER IX

SUGGESTIONS FOR FUTURE EXPERIMENTAL WORK

Three aspects of the brominative decarboxylation are in definite need of further clarification by experimental means. One is a determination of whether or not the hydroxybenzoic acid loses a hydrogen ion prior to, simultaneously with, or after the initial attack by the bromine molecule. As has been pointed out, the first possibility can be distinguished from the latter two at zero bromide ion concentration and varying hydrogen ion concentration. Since under these conditions the first step is irreversible and rate controlling, the hydrogen ion dependence of the first step only would be determined.

The second problem is that of determining more precisely the rate of bromination of dibromo-p-hydroxybenzoic acid at very low bromide ion concentrations. If it can be shown that as the bromide ion concentration approaches zero, $1/V$ approaches a very small number (approximately zero), the mechanism involving Br^+ (or H_2OBr^+) is confirmed.

The third problem is to determine if carbon dioxide is indeed eliminated in the rate determining step and under what conditions. According to the proposed mechanism for dibromosalicylic acid, at high bromide ion concentrations

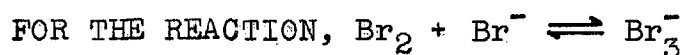
carbon dioxide- C^{14} should be lost more slowly than carbon dioxide- C^{12} ; while at low bromide ion concentrations, the first step of the reaction becomes essentially irreversible and no isotope effect should be observed. If an isotope effect that is dependent on the bromide ion concentration is observed for dibromo-p-hydroxybenzoic acid, the conclusion would be justified that the para compound is brominated by the same mechanism as the ortho compound. On the other hand, if no isotope effect is observed or an isotope effect that is independent on the bromide ion concentration is observed for dibromo-p-hydroxybenzoic acid, the conclusion would be that bromination is by Br^+ or H_2OBr^+ .

Since it has been demonstrated in other brominations that they are catalyzed by salts of weak acids due to the formation of acyl-hypobromites which are capable of functioning as brominating agents (36, 39, 34, 41), it is suggested that a detailed study of specific salt effects might be in order. This might demonstrate whether or not nucleophiles other than bromide ion could cause reversibility of the first step. It is very probable that no such effects would be noted in media of low pH; however, as the hydrogen ion concentration was made lower and the pH approached 7 or more, such effects might become pronounced.

APPENDICES

APPENDIX A

THE DETERMINATION OF THE EQUILIBRIUM CONSTANT



The equilibrium constant for the reaction was determined in 70, 75, and 80 per cent acetic acid by the method described by Popov and co-workers (83) using a Beckman, Model DU, spectrophotometer thermostated at 20°. Calibrated Corex cells of optical path length 1.000 ± 0.001 cm. and equipped with ground stoppers were used. The solvents used were prepared according to the procedure described for Runs 54 and following.

The procedure is based on the assumption that the total optical density of a solution at a given wave length is equal to the sum of the optical densities of the component parts of the solution. In the system being studied and at the wave lengths being observed, the only colored species present are the tribromide ion and free bromine. This system then would be represented by the equation

$$D_{\text{tot}}^{\lambda} = e_{\text{Br}_2}^{\lambda} (\text{Br}_2) + e_{\text{Br}_3^-}^{\lambda} (\text{Br}_3^-) .$$

83. A. I. Popov, K. C. Brinker, L. Companaro, and R. W. Rinehart, J. Am. Chem. Soc., **73**, 514 (1951).

The extinction coefficient for free bromine, $\epsilon_{\text{Br}_2}^\lambda$, can be determined from the optical density, D , of a solution of bromine of known concentration and containing no bromide ion, and the extinction coefficient for the tribromide ion, $\epsilon_{\text{Br}_3^-}^\lambda$, from the optical density of a solution of bromine containing so much bromide ion that all of the bromine has been converted to tribromide ion. Hence, if the total optical density of solutions containing both free bromine and tribromide ion are known, and the extinction coefficients for each species are also known, it is possible to solve for the concentration of each species, and from them calculate the equilibrium constant.

By surveying the spectrum of various mixtures of bromine and bromide ion in acetic acid from 250-500 millimicrons, a peak was found in the spectrum of bromine at about 400 millimicrons and in the spectrum of tribromide ion at about 250 millimicrons. The tribromide extinction coefficient was very much greater than that of bromine, but it was decided that if the tribromide ion concentration was kept fairly low, satisfactory measurements could be made at 360, 380, 400, and 420 millimicrons. Ideally, for accurate determinations of concentrations of species, concentrations and extinction coefficients of the contributing species should be of approximately the same order of magnitude. This was not possible in this system and the

accuracy of the determinations suffered accordingly.

The extinction coefficient of tribromide ion had to be determined by approximation. If a solution were made 0.300 M in bromide ion and 0.005 M in bromine almost all of the bromine would be tied up as tribromide ion but there would still be a small contribution from free bromine. If it is assumed that the optical density is due entirely to tribromide ion and is used to calculate an approximate equilibrium constant, this approximate value of the equilibrium constant may then be used to re-evaluate the extinction coefficient for tribromide. It was found that the correction was so small that only one correction was sufficient.

Some sample calculations will be shown to illustrate the method used, following which a tabulation of results will be shown.

1. Determination of $e_{\text{Br}_2}^{400}$, in 80 per cent acetic acid, 0.300 M in HClO_4 , and 0.00568 M in Br_2 .

$$D = 1.000$$

$$D = e_{\text{Br}_2}(\text{Br}_2)$$

$$1.000 = e_{\text{Br}_2}(0.00568)$$

$$e_{\text{Br}_2} = 176.1$$

2. Determination of $e_{\text{Br}_3}^{400}$, in 80 per cent acetic acid, 0.300 M in HBr , and 0.00118 M in Br_2 .

$$D = 0.642$$

$$D = e_{\text{Br}_3^-}(\text{Br}_3^-)$$

$$0.642 = e_{\text{Br}_3^-}(0.00118)$$

$$e_{\text{Br}_3^-} = 544.1$$

By calculations such as these, extinction coefficients were determined in 70, 75, and 80 per cent acetic acid at 360, 380, 400, and 420 millimicrons.

It is readily seen that there is little if any solvent dependency in this system, as far as the extinction coefficients are concerned.

In Table 82 are shown the data obtained and the equilibrium constants calculated from them, after re-evaluating the extinction coefficient for tribromide ion.

Table 81

Average Extinction Coefficients of Br_2 and Br_3^- 1. e_{Br_2}

Solvent (per cent HOAc)	360 mu.	380 mu.	400 mu.	420 mu.	Number averaged
70	108	166	178	157	3
75	107	166	180	160	2
80	105	164	179	159	4

2. $e_{\text{Br}_3^-}$ (initial uncorrected values)

70	936	736	545	307	2
75	940	739	548	308	2
80	938	738	548	310	2

3. $e_{\text{Br}_3^-}$ (corrected values)

70	967	757	556	309	
75	961	751	555	308	
80	970	757	557	311	

Table 82

Evaluation of Equilibrium Constants from
Observed Optical Densities

1. 80.0 per cent Acetic Acid, 0.00364 M in Br₂, 0.00100 M in HBr and 0.299 M in HClO₄
2. 80.0 per cent Acetic Acid, 0.00448 M in Br₂, 0.00150 M in HBr and 0.2985 M in HClO₄
3. 75.0 per cent Acetic Acid, 0.00373 M in Br₂, 0.00100 M in HBr and 0.299 M in HClO₄
4. 75.0 per cent Acetic Acid, 0.00508 M in Br₂, 0.00150 M in HBr and 0.2985 M in HClO₄
5. 70.0 per cent Acetic Acid, 0.00404 M in Br₂, 0.00100 M in HBr and 0.299 M in HClO₄
6. 70.0 per cent Acetic Acid, 0.00416 M in Br₂, 0.00150 M in HBr and 0.2985 M in HClO₄

#	Observed Optical Densities				Average K
	360 mu.	380 mu.	400 mu.	420 mu.	
1	0.596	0.738	0.741	0.616	
2	0.842	0.981	0.955	0.772	93 ± 3
3	0.619	0.763	0.756	0.625	
4	0.928	1.098	1.067	0.869	85 ± 9 ^a
5	0.649	0.820	0.819	0.680	
6	0.751	0.897	0.877	0.710	83 ± 3 ^b

^aIt is of interest to compare this data with the equilibrium constant obtained by Bradfield, Davies, and Long (38) in 75 per cent acetic acid and at ionic strengths of ca. 0.01-0.02. Using an aspiration technique, they obtained a value of 91 ± 3.

^bThe average value for 70 per cent acetic acid is the average of values obtained from line 6 only.

APPENDIX B

SPECIMEN PREPARATIONS OF ACETIC ACID SOLVENTS

Solvents Used in Runs 18-47

In preparing these solvents, all components were mixed in proportions so that the ratio of acetic acid to water by weight would be the desired amount. It was necessary therefore to know the amount of water in all of the reagents. As previously described, the amount of water in glacial acetic acid was determined by Karl Fischer titration. In the case of the other ingredients, molar concentration was converted to per cent composition and at the same time the density determined by use of the tables of per cent composition versus density of aqueous solutions found in the International Critical Tables (79). All solutions were measured at 20°, using the minimum number of pipets, allowing them to drain for 10 seconds and shaking each 10 times. The approximate loss in volume in mixing was determined in advance by separate experiments in order to calculate the amount of reagent necessary to use to get the desired bromide and hydrogen ion concentrations. The actual final volume of the mixture was ascertained by mixing the solutions in a volumetric flask having a roughly calibrated

neck.

By accounting for the possible errors in densities and pipet volumes, it was found that the solvents would be within 0.2 per cent of the calculated value. Unfortunately, when the solvents were actually prepared for Runs 18-49, the densities of the lithium bromide solutions used were assumed to be the same as water. When the densities were later calculated and the concentrations of the solvents corrected, this error was found to have made the solvents 0.2 per cent less in water than had been desired. Sample preparations of two of these solvents will be shown.

75 per cent acetic acid.--The following reagents were used:

Acetic acid, 0.88% water	- 1.052 g./ml.
Water	- 0.998 g./ml.
70.62% Perchloric acid, 11.78 M	- 1.677 g./ml.

Start with 750 ml. of acetic acid:

750 ml. x 1.052 g./ml.	= 789 g.
789 g. x 0.0088	= 6.94 g. water
Actual acetic acid	= 782.06 g.

Total weight of solution based on 782.06 g. of acetic acid:

0.75x = 782.06 g.
x = 1042.75 g.

Total weight of water required:

1042.75 g. of solution
- 782.06 g. of acetic acid
260.69 g. of water

Volume of 70% perchloric acid required to make a 0.1 N solution. (Approximate volume of solution will be 978 ml.)

$$V \times 11.78 = 978 \times 0.1$$

$$V = 8.30 \text{ ml.}$$

Weight of water in 8.30 ml. of 70% perchloric acid:

$$8.30 \text{ ml.} \times 1.677 \text{ g./ml.} \times 0.294 = 4.10 \text{ g.}$$

Volume of water to be added to solution:

Total water necessary	260.69 g.
Water in acetic acid	-6.94 g.
Water in perchloric acid	<u>-4.10 g.</u>
Water to be added	249.65 g.

$$249.65 \text{ g.} + 0.998 \text{ g./ml.} = 250.15 \text{ ml.}$$

80 per cent acetic acid, 0.1 N in lithium bromide.--The following reagents were used:

Acetic acid, 0.88%	- 1.052 g./ml.
Water	- 0.998 g./ml.
70.62% Perchloric acid, 11.78 M	- 1.677 g./ml.
Lithium bromide, 0.5107 M, 4.30%	- 1.030 g./ml.

Start with 400 ml. of acetic acid:

$$400 \text{ ml.} \times 1.052 \text{ g./ml.} = 420.80 \text{ g.}$$

$$420.8 \text{ g.} \times 0.0088 = \underline{3.71} \text{ g. water}$$

$$\text{Actual acetic acid} = 417.09 \text{ g.}$$

Add 98.7 ml. of lithium bromide solution (the volume was that calculated to be necessary using pure water):

$$98.7 \text{ ml.} \times 1.030 \text{ g./ml.} \times 0.9570 =$$

$$97.28 \text{ g. water}$$

Volume of 70% perchloric acid required to make a 0.1 N solution. (Approximate volume of solution will be 484.1 ml.)

$$V \times 11.78 = 484.1 \times 0.1$$

$$V = 4.11 \text{ ml.}$$

Weight of water in 4.11 ml. of 70% perchloric acid:

$$4.11 \text{ ml.} \times 1.677 \text{ g./ml.} \times 0.294 = 2.06 \text{ g.}$$

Actual water added:

From lithium bromide solution	97.28 g.
From acetic acid	3.71 g.
From perchloric acid	<u>2.06</u> g.
Total water	103.05

Actual solvent concentration:

$$\frac{417.09 \text{ g.}}{417.09 + 103.05} \times 100 = 80.19\% \text{ acetic acid}$$

Solvents Used in Runs 48-53

The solvents used in these runs were prepared by the same technique as previously described, with one notable exception--the per cent composition and density of the hydrogen bromide solution was determined and used in the calculations. A sample calculation for the preparation of one of these solvents is shown.

80 per cent acetic acid, 0.2 N in hydrogen bromide.--The following reagents were used:

Acetic acid, 0.39% water	1.0508 g./ml.
7.89% hydrogen bromide, 1.0306 M	1.0561 g./ml.

Start with 400 ml. of acetic acid:

$$\begin{aligned} 400 \text{ ml.} \times 1.0508 \text{ g./ml.} &= 420.32 \text{ g.} \\ 420.32 \text{ g.} \times 0.0039 &= \underline{1.64 \text{ g.}} \text{ water} \\ \text{Actual acetic acid} &= 418.68 \text{ g.} \end{aligned}$$

Total weight of solution based on 418.68 g. of acetic acid;

$$\begin{aligned} 0.80x &= 418.68 \text{ g. of acetic acid} \\ x &= 523.35 \text{ g. of solution} \end{aligned}$$

$$\begin{aligned} &- 418.68 \text{ g. of acetic acid} \\ &\underline{104.67 \text{ g. of water required}} \end{aligned}$$

$$- \frac{1.64 \text{ g. water in acetic acid}}{103.03 \text{ g. water yet to be added}}$$

Volume of 7.89% hydrogen bromide necessary to supply 103.03 g. of water:

$$\begin{aligned} 0.9211x &= 103.03 \text{ g. of water} \\ x &= 111.85 \text{ g. of hydrogen bromide} \\ &\quad \text{solution} \end{aligned}$$

$$111.85 \text{ g.} \div 1.0561 = 105.91 \text{ ml.}$$

Total volume of solution was 491.5 ml. Concentration of hydrogen bromide then was:

$$N = (105.91 \text{ ml.})(1.0306 \text{ N}) / (491.5 \text{ ml.}) = 0.2221$$

Solvents Used in Runs 54-83

These solvents were the most accurately prepared and the most reproducible of all the solvents used. In addition to determining the densities and per cent composition of all the components (as previously described), a further refinement was used which made it possible to determine in advance the final concentration of the additions, whatever the loss in volume on mixing the various ingredients might be. This was done merely by calculating the desired quantities of the solutions on the basis of a final volume of 500 ml. Since on mixing, the total volume of solution was always less than 500 ml., all that was necessary was to add enough of a water-acetic acid mixture of the desired composition to bring the liquid level to the mark in a 500 ml. volumetric flask. Two sample calculations are shown.

80.0 per cent acetic acid, 0.100 N in LiBr, 0.200 N in HBr.--
The following reagents were used:

Acetic acid, 0.62% water	1.0514 g./ml.
Water	0.9982 g./ml.
12.56% Lithium bromide, 1.584 N	1.0957 g./ml.
11.27% Hydrogen bromide, 1.508 N	1.0825 g./ml.

Start with 400 ml. of acetic acid:

$$\begin{aligned}
 400 \text{ ml.} \times 1.0514 \text{ g./ml.} &= 420.56 \text{ g.} \\
 420.56 \text{ g.} \times 0.0062 &= 2.61 \text{ g. water} \\
 \text{Actual acetic acid} &= 417.95 \text{ g.}
 \end{aligned}$$

Total weight of solution based on 417.95 g. of acetic acid:

$$\begin{aligned}
 0.8x &= 417.95 \text{ g. acetic acid} \\
 x &= 522.44 \text{ g. solution}
 \end{aligned}$$

Total weight of water necessary:

$$\begin{aligned}
 &522.44 \text{ g. of solution} \\
 &\underline{-417.95 \text{ g. of acetic acid}} \\
 &104.49 \text{ g. of water}
 \end{aligned}$$

Hydrogen bromide solution necessary to make a solvent 0.200 N:

$$\begin{aligned}
 V \times 1.508 &= 500 \times 0.2 \\
 V &= 66.3 \text{ ml.}
 \end{aligned}$$

$$\begin{aligned}
 66.3 \text{ ml.} \times 1.0825 \text{ g./ml.} \times 0.8873 &= \\
 63.68 \text{ g. water}
 \end{aligned}$$

Lithium bromide solution necessary to make a solvent 0.100 N:

$$\begin{aligned}
 V \times 1.584 &= 500 \times 0.1 \\
 V &= 31.6 \text{ ml.}
 \end{aligned}$$

$$31.6 \text{ ml.} \times 1.0957 \text{ g./ml.} \times 0.8744 = 34.62 \text{ g. water}$$

Remaining water to be added:

Total water necessary	104.49 g.
Water in acetic acid	- 2.61 g.
Water in hydrogen bromide soln.	-63.68 g.
Water in lithium bromide soln.	-30.27 g.
Water necessary	<u>7.93 g.</u>

$$7.93 \text{ g.} \div 0.9982 \text{ g./ml.} = 7.94 \text{ ml.}$$

Add enough 80 per cent acetic acid to make a total volume of 500 ml.

75 per cent acetic acid, 0.100 N in HBr, 0.200 N in HClO_4 .--
The following reagents were used:

99.41% Acetic acid	1.0513 g./ml.
Water	0.9982 g./ml.
11.63% Hydrogen bromide, 1.560 N	1.0853 g./ml.
70.42% Perchloric acid, 11.73 N	1.674 g./ml.

Start with 375 ml. of acetic acid:

$$\begin{aligned} 375 \text{ ml.} \times 1.0513 \text{ g./ml.} &= 394.24 \text{ g.} \\ 394.24 \text{ g.} \times 0.0059 &= 2.33 \text{ g. water} \\ \text{Actual acetic acid} &= 391.91 \text{ g.} \end{aligned}$$

Total weight of solution based on 391.91 g. of acetic acid:

$$\begin{aligned} 0.75x &= 391.91 \text{ g. acetic acid} \\ x &= 522.55 \text{ g.} \end{aligned}$$

Total weight of water necessary:

$$\begin{aligned} &522.55 \text{ g. of solution} \\ &\underline{-391.91 \text{ g. of acetic acid}} \\ &130.64 \text{ g. of water.} \end{aligned}$$

Hydrogen bromide solution necessary to make a solvent 0.100 N in HBr:

$$\begin{aligned} V \times 1.560 &= 500 \times 0.1 \\ V &= 32.05 \text{ ml.} \end{aligned}$$

$$32.05 \text{ ml.} \times 1.0853 \text{ g./ml.} \times 0.8837 = 30.74 \text{ g. water}$$

Perchloric acid necessary to make a solvent 0.200 N in HClO_4 :

$$\begin{aligned} V \times 11.73 &= 500 \times 0.2 \\ V &= 8.52 \text{ ml.} \end{aligned}$$

$$8.52 \text{ ml.} \times 1.674 \text{ g./ml.} \times 0.2958 = 4.22 \text{ g. water}$$

Remaining water to be added:

Total water necessary	130.64 g.
Water in acetic acid	- 2.33 g.
Water in hydrogen bromide soln.	-30.74 g.
Water in perchloric acid	- 4.22 g.
Water necessary	<u>93.35 g.</u>

$93.35 \text{ g.} \div 0.9982 \text{ g./ml.} = 93.52 \text{ ml.}$

Add enough 75 per cent acetic acid to make a total volume of 500 ml.

APPENDIX C

SAMPLE CALCULATION OF INTEGRATED RATE CONSTANTS

The rate constants were calculated from the integrated form of the second order rate equation, i.e.,

$$k = \frac{2.303}{(b-a)t} \log \frac{a(b-x)}{b(a-x)},$$

in which k is the rate constant (l./mol.min.), t is the time elapsed for any point, a is the initial stoichiometric concentration of DBPHBA, DBSA, or DBP, b is the initial stoichiometric concentration of bromine, and x is the molar concentration of substance reacted (or product formed) at time t. The sample calculation will be on Run 55, Table 50.

a = 0.004458 M a/b = 0.5475 (Na₂S₂O₃) = 0.02190 N
b = 0.008142 M b-a = 0.003684 M (50 ml. samples)

t (min)	Titer (ml.)	(b-x)	(a-x)	(x)	$\frac{(b-x)}{(a-x)}$
476	34.64	0.007586	0.003902	0.000556	1.944
1281	31.15	0.006822	0.003138	0.001320	2.174
1895	29.28	0.006427	0.002743	0.001715	2.343
2916	26.66	0.005852	0.002168	0.002290	2.699
4593	23.94	0.005255	0.001571	0.002887	3.345
7735	21.03	0.004616	0.000932	0.003526	4.953

$$k = \frac{2.303}{(0.003684)t} \log (0.5475) \frac{(b-x)}{(a-x)}$$

t (min.)	$\frac{2.303}{(b-a)t}$	$\frac{a(b-x)}{b(a-x)}$	$\log \frac{a(b-x)}{b(a-x)}$	$\frac{2.303}{(b-a)t} \log \frac{a(b-x)}{b(a-x)}$
476	1.313	1.0643	0.02706	0.0355
1281	0.4880	1.1903	0.07566	0.0369
1895	0.3299	1.2828	0.10816	0.0357
2916	0.2144	1.4777	0.16958	0.0364
4593	0.1361	1.8314	0.26279	0.0358
7735	0.08081	2.712	0.43329	0.0350

All calculations of this type were made on an electric calculating machine with use of the number of significant figures shown.

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BIOGRAPHICAL SKETCH

Ulysses Virgil Henderson, Jr., son of U. V. Henderson and Mildred Gibbons Henderson, was born July 15, 1926, in Miami, Florida. In 1930 the family moved to West Palm Beach, Florida, where he attended elementary and high schools. He was graduated from Palm Beach High School with the Class of 1943.

In June, 1943, he enrolled in the Georgia School of Technology where he began his study of chemistry. Due to the war, his studies were interrupted and in November, 1944, he commenced basic training in the U. S. Merchant Marine Cadet Corps at Pass Christian, Mississippi. The year between April, 1945, and May, 1946, was spent in training at sea aboard the S. S. Andrew Jackson and the S. S. Junior. From June, 1946, until December, 1947, he was stationed at the U. S. Merchant Marine Academy, Kings Point, New York, from which he was graduated with a Third Mate's License, a commission as Ensign, U. S. N. R., and a B. S. degree. He sailed as Jr. Third Officer of the S. S. American Clipper for the United States Lines Company from January until September of 1948.

He returned to the Georgia Institute of Technology in September, 1948, and completed his work for the degree of B. S. (with honor) in Chemistry in September, 1950.

Immediately thereafter, he entered graduate school at the same institution, serving as a Graduate Assistant for three years and holding a Graduate Division Fellowship for one year.

In September, 1951, he married Evelyn Foster of Atlanta and McDonough, Georgia.

Upon completing work for his doctorate, he was employed by the Standard Oil Company of Indiana at their Whiting Research Laboratory in Whiting, Indiana.