KINETIC STUDIES OF FAST BROMINATION REACTIONS
WITH A CONSTANT CURRENT GENERATOR AND
AMPEROMETRIC INDICATOR SYSTEM

by

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A Dissertation Submitted to the Faculty of the
DEPARTMENT OF CHEMISTRY

In Partial Fulfillment of the Requirements

For the Degree of

DOCTOR OF PHILOSOPHY

In the Graduate College

THE UNIVERSITY OF ARIZONA

1966

THE UNIVERSITY OF ARIZONA

GRADUATE COLLEGE

I hereby recommend that this dissertation prepared under my
direction by George W. O'Dom
entitled Kinetic Studies of Fast Bromination Reactions with a
Constant Current Generator and Amperometric Indicator
System.
be accepted as fulfilling the dissertation requirement of the
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FORWARD

The author wishes to express his sincere appreciation to Dr. Quintus Fernando, who directed this work, for his constant encouragement, and many worthwhile suggestions in both the experimental work and in the preparation of this dissertation.

The author also wishes to thank Dr. Henry Freiser for the interest and helpful suggestions provided throughout the course of this work.

The author is especially grateful to his wife, Eleanor, whose years of sacrifice and work helped make this possible.

Financial assistance from the National Institute of Health is gratefully acknowledged.

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Abstract

TITLE: KINETIC STUDIES OF FAST BROMINATION REACTIONS WITH
A CONSTANT CURRENT GENERATOR AND AMPEROMETRIC
INDICATOR SYSTEM

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Ph.D.

The University of Arizona

1966

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Part I

A commercially available constant current source and a recording polarograph are used to measure the kinetics of several bromination reactions in aqueous solutions. The effect of varying the rate of stirring and concentration of reactants on the rate constants is determined. The method is confirmed by comparing the rate constants obtained for the bromination of acetone and m-nitrophenol with values found in the literature. An observed second-order rate constant of $10^5 \text{ M}^{-1} \text{ sec}^{-1}$ was found to be an upper limit for this method.

The method is further investigated by a detailed rate study of the monobromination of l-tyrosine and related compounds. The importance of the shape of the current-time curves to kinetic studies by this method and the effect of current damping on these curves are discussed. In compounds which undergo monobromination

a decrease in the observed rate constant with time is observed for calculations made in the region where the slope of the curve is large. The monobromination of tyrosine is a second-order reaction. The rate constant is independent of hydrogen ion concentrations up to the region of $[H^+] = K_{a1}$, since the bromination rate constants for the neutral and protonated forms are the same. Furthermore, tribromide ion is not an active brominating species in these reactions. Ortho-tyrosine brominates with a rate constant which is about five times larger than that of tyrosine and about the same as that of phenol.

Part II

The dibromination kinetics of 8-quinolinol and 4-methyl-8-quinolinol are studied. The first and second bromination rate constants of these compounds are separated by simulation of the concentration-time data on an analog computer. The dibromination is assumed to occur by a consecutive parallel mechanism. A good fit between the computer readout and the experimental data confirms the mechanism.

The monobromination of 8-quinolinol and 4-methyl-8-quinolinol have bromination rate constants which are about the same. The rate constant for the neutral species is about 10^4 times greater than that of the protonated

I. INTRODUCTION

Constant current coulometry was first applied to the problem of chemical kinetics in 1954. In a short note Pearson and Piette¹ described the use of constant current coulometry to obtain the overall rate constant of the reaction

In this study a relatively large concentration of nitroethane was used and hydroxide ions were generated electrolytically. The method is dependent upon achieving a steady state where the rate of generation is equal to the rate of the chemical reaction. In the presence of a constant concentration of hydroxide ion, pseudo first-order kinetics were obtained. Furthermore, a second-order rate constant of $3 \times 10^2 \text{M}^{-1} \text{sec}^{-1}$ was calculated from a knowledge of the steady state hydroxide ion concentration. The authors claimed that this method was useful for the calculation of rate constants of relatively fast reactions. Shortly thereafter a complete study of the coulometric method was published by other workers²

¹Pearson, R. G. and Piette, L. T., J. Amer. Chem. Soc. 76, 3087 (1954)

²Farrington, P. S. and Sawyer, D. T., J. Amer. Chem. Soc. 78, 5536 (1956)

who used the method to study the basic hydrolysis of some esters with second-order rate constants of $1-15 \text{ M}^{-1}\text{sec}^{-1}$. They also extended the method to the hydrogenation of olefins in non-aqueous solvents. At about the same time an instrument which automatically adjusted the current to the desired steady state value was described in the literature¹. This instrument was used for the measurement of enzyme activity in systems in which hydrogen ions were produced. As the enzyme reaction proceeded, hydrogen ions were formed; the circuit became unbalanced; and hydrogen ions were removed by electrolysis to restore the balance. By this method a steady state was achieved where the rate of reaction was related to the current.

A variation of constant current coulometry has been applied to the study of the kinetics of redox reactions². Consider a reaction of the type

During the course of the reaction the ratio of M^{++}/M^+ decreases and the potential given by the $\text{M}^{++}-\text{M}^+$ couple decreases. If the ratio M^{++}/M^+ is maintained constant by generation of M^{++} from M^+ , the current necessary to keep

¹Einsel, Jr., D.W., Trurnit, H. J., Silver, S. D., and Steiner, E. C., Anal. Chem. 28, 408 (1956)

²Matsen, J. M., Linford, H. B., Anal. Chem. 34, 142-5 (1962)

this ratio constant is related to the rate of the reaction. An apparatus is described which is effective for slow reactions (i.e. currents of 1-60 milliamps).

If, in the above reaction, one starts with the constant potential electrolysis of M^+ , the current will decay to a certain value and then level off. This phenomenon has been observed in another system¹, and the steady state current was called a "kinetic" current. Recently this method was applied to kinetic studies². The reason for the initial decay is the depletion of the M^+ species. This decrease in M^+ will cause the chemical reaction rate to increase due to an increase in M^{++} . Under steady state conditions the increase in M^+ from the chemical reaction is balanced by the decrease in M^+ due to electrolysis. It was estimated that rate constants as large as $10^5 M^{-1} \text{sec}^{-1}$ could be measured for this reaction, although the method was actually used to calculate a rate constant of $3.2 \pm .1 \times 10^{-4} M^{-1} \text{sec}^{-1}$.

Meites et. al.³ derived kinetic equations for automatic and coulometric titration curves in which the

¹Rechnitz, G. A. and Laitinen, H. A., Anal. Chem. 36, 2265 (1964)

²Rechnitz, G. A. and McClure, J. E., Talanta 12, 153 (1965)

³Cover, R. E. and Meites, L., J. Phys. Chem. 67, 1528, 2311 (1963)

reactant is added without interruption. The equations were derived for systems in which complex formation, disproportionation, and catalytic competing reactions preceeded an irreversible redox reaction step.

One area which has not been discussed as yet and which has received much attention is kinetic studies with electro-generated bromine. The bromine-bromide couple is reversible, and can be generated with 100% current efficiency. This fact was utilized by Dubois and coworkers¹ to devise an instrument for the generation of bromine for kinetic studies. Bromine was introduced by electrolysis into the system where it was free to react with a substance in solution. Dual polarized electrodes were used to detect bromine down to $1 \times 10^{-7} \text{ M}$ and the concentration of bromine was maintained at a constant value by electrolytic generation. In this manner pseudo first-order kinetics were obtained for second-order reactions, and from the results obtained, the second-order rate constants were calculated. Rate constants as large as $1.7 \times 10^6 \text{ M}^{-1} \text{ sec}^{-1}$ were reported. This method was used to study the bromination of olefins² and the bromination of phenolic ethers such as anisole³.

¹Dubois, J.E. and Mouirer, G., Compt. Rend. 257, 1104 (1962)

²Dubois, J. E. and Schwarcz, Compt. Rend. 259, 2227 (1964)

³Dubois, J.E. and Aaron, J. J., Compt. Rend. 258, 2313 (1964)

Fernando et.al.¹ have described a pulse generator which will generate bromine with 100% current efficiency. This apparatus when coupled with a single polarized electrode amperometric detection system was used to study the monobromination of phenol², phenetole³, and 8-quinolinols⁴. Although Dubois claims to have developed the constant current coulometry-amperometric indicator system⁵ no experimental details have been reported.

Although a second-order rate constant may be quite large, say $10^6 \text{ M}^{-1} \text{ sec}^{-1}$, it is the product of this rate constant and the concentrations of the reactants which determines the overall rate of change of concentration with respect to time. Therefore if the concentrations of reactants are kept small, the rate can be brought into a measurable region⁶. These concentrations may be controlled

¹Fernando, Q. and Devanathan, M.A.V., Trans.Faraday Soc. 52, 1332 (1956)

²Kozak, G. and Fernando, Q., Anal. Chim. Acta. 26, 541-47 (1962)

³Kozak, G. and Fernando, Q., J. Phys. Chem. 67, 811 (1963)

⁴Kozak, G. and Fernando, Q., and Freiser, H., Anal. Chem. 36, 296-300 (1964)

⁵Reference to work to be published in Compt. Rend. 258, 2313 (1964)

⁶Caldin, E. F., Fast Reactions in Solution, Wiley 1964

by physical or chemical methods. The chemical control can be achieved if equilibrium is established between a less reactive form and reactive form of the compound. For example, the equilibrium between tribromide ion and bromine regulates the concentration of the more reactive species, bromine. Furthermore, if the species to be brominated is a base such as an amine, then the use of high acidity will insure that most of the compound is present in the usually less reactive acid form.

An example of chemical control is found in the study of the bromination of aromatic amines¹ which was carried out in 1-8M sulfuric acid. With bromine concentrations in the range of $3 \times 10^{-3} \text{M}$, rate constants ranging from $10^6 \text{M}^{-1} \text{sec}^{-1}$ to $10^{10} \text{M}^{-1} \text{sec}^{-1}$ were measured. The rate constant for the bromination of acetone was obtained² by taking advantage of the fact that only the enol form will brominate. Thus, because of a keto-enol equilibrium constant of 2×10^{-6} , the concentration of enol is very small and a rate constant for the bromination of acetone of $5 \times 10^6 \text{M}^{-1} \text{sec}^{-1}$ was calculated. Moreover the rate constant for the bromination of phenol was obtained in aqueous

¹Bell, R. P. and Ramsden, D., J. Chem. Soc. 1958, 161

²Bell, R. P. and Davis, G. G., J. Chem. Soc. 1964, 902

acid¹ by using low concentrations of both phenol and bromine. Therefore, just from a few selected examples one can see that the use of low concentrations is an effective means of studying fast bromination reactions.

The use of low concentrations of reactants, however, places a burden on the indicator system. In most cases bromine concentrations used are much too low to be measured by classical methods, and physical methods must be used. As mentioned, Dubois has used amperometry with dual polarized electrodes to measure bromine down to about $1 \times 10^{-7} \text{M}$.² Bell and coworker used a potentiometric method to measure bromine concentration³. It was found that the potential of the bromine-bromide couple varied linearly with time down to bromine concentrations of 10^{-6} to 10^{-7}M . In a later paper a concentration range of 10^{-8}M to 10^{-9}M is quoted as marking a lower limit for bromine measurement⁴. This value seems somewhat low for

¹Kozak, G. and Fernando, Q., Anal Chim. Acta. 26, 541-47 (1962)

²Dubois, J. E. and Mouvier, G., Tetrahedron Letters 1963, 1325; Dubois, J. E. and Schwarcz, Compt. Rend. 259, 2227 (1964); Dubois, J. E., Z. fur Elektrochem. 64, 143 (1960); Dubois, J. E. and Mouvier, G., Compt. Rend. 257, 1104 (1962)

³Bell, R. P. and Ramsden, E., J. Chem. Soc. 1958, 161

⁴Bell, R. P. and Davis, G. G., J. Chem. Soc. 1964, 902

potentiometric measurements in view of the possibility of obtaining mixed potentials at the electrodes. The rotating platinum microelectrode has been used effectively to measure bromine in the region of $10^{-7}M$.¹

The manner in which the experimental results are treated to obtain rate constants varies widely from one worker to another. The constant concentration method outlined above greatly simplifies the calculation, for the reactions are pseudo first-order. The potentiometric method also gives pseudo first-order kinetics when a relatively large concentration of substrate is present. In this case the potential vs time plots are linear and the rate of the reaction is related to the potential-time slope given by equation I-1, in which dE/dt is known and therefore k_1 can be calculated. A method based on the analysis of current-time curves was introduced by

$$k_1 = \frac{d \ln [Br_2]}{dt} = 0.0782 \frac{dE}{dt} \quad \text{I-1}$$

Fernando et.al.² This method, which will be outlined in detail later, makes it possible to calculate the rate of

¹Kozak, G. and Fernando, Q. Anal. Chem. Acta. 26, 541 (1962); Kozak, G. and Fernando, Q. J. Phys. Chem. 67, 811 (1963); Kozak, G., Fernando, Q., and Freiser, H., Anal. Chim. 36, 296 (1964)

²Kozak, G. and Fernando, Q., J. Phys. Chem. 67, 811 (1963)

consumption of bromine at any point on the curve by a glass rod technique^{1,2}. Using this, one can calculate the rate constant for any simple second-order reaction where the concentration of both reactants are varying. This method therefore allows one to get away from steady-state conditions altogether.

Recently a paper appeared whose stated purpose was to simplify the calculations used by Kozak and Fernando. In this method³ the integrated rather than the differential forms of the kinetic equations were used. The instantaneous concentrations of reactants are calculated essentially in the same manner as the Fernando method with the exception that the concentration of the substrate at time t is obtained by turning the generator off and observing the decrease in the indicator current. This procedure is at best risky, since for a fast reaction the response speed of the recorder cannot be neglected.

As can be seen constant current coulometry is a valuable method for the study of chemical kinetics. Moreover, it is especially useful for the study of fast bromination reactions. For this reason, the present study

¹Hoare, J., J. Chem. Educ. 38, 570 (1961)

²Hockandel, C. J., Ibid 39, 299 (1962)

³Janata, J. and Zyka, J., J. Collection Czech. Chem. Commun. 30, 1723 (1965)

was undertaken to establish the useful range over which the apparatus used in this study is effective for fast bromination kinetic studies. This apparatus is commercially available and differs from that used by Kozak¹ in that a constant current source is used in place of a pulse generator. Furthermore, it seemed desirable to investigate the kinetics of a system which undergoes an uncomplicated monobromination reaction in order to further establish the method.

¹Kozak, G., Ph.D. Thesis, University of Arizona
(1963)

I-II. EXPERIMENTAL

A. Apparatus

A commercially available Sargent Coulometric Current Source Model IV was used to provide a constant current for bromine generation. This source is capable of delivering currents of 5 to 200 milliamps which are constant to 0.1%. The source current of the two lowest scales was calibrated by measuring the potential drop across a 9.559 ohm standard resistance with a Minneapolis-Honeywell Model 2730 potentiometer. Current values of 4.830 milliamps and 9.662 milliamps were found for the instrument settings of 0.005 and 0.01 respectively. The listed values for these settings are 4.825 milliamps $\pm$ 0.1% and 9.66 milliamps. The current was checked for constancy with varying circuit resistance by inserting a variable resistance in series with the standard resistance. This resistance was varied from zero to about 5k ohms and no change in the potential drop across the standard resistance was observed. The polarizing voltage and the recorder for the amperometric indicator system was provided by a Sargent Model XV Polarograph. A mercury sealed platinum microelectrode, rotated at 600 r.p.m., in conjunction with a standard calomel reference electrode comprised the indicator system. The micro-electrode was

polarized to 0.4V vs S.C.E. The generator electrodes at first consisted of a platinum foil electrode (6.5cm^2) connected by means of a side arm to a saturated calomel electrode. It was found, however, that after several runs had been made, the calomel electrode showed separation between the mercury and the calomel. The calomel, therefore, had to be replaced frequently. A platinum foil electrode was tried in an acid solution in place of the calomel electrode. During bromine generation, however, this electrode became covered with hydrogen, and the bromine generation was affected. Finally, a silver foil electrode (40cm^2) immersed in 0.1M perchloric acid was found to be satisfactory for use as an auxiliary electrode in the generating system.

The electrolysis cell consisted of a 100 ml. central compartment connected by two side arms to the calomel and the silver foil electrodes. Medium porosity sintered glass frits and agar plugs saturated with sodium nitrate separated the central compartment from the side arms. The whole cell was arranged in a triangular pattern with the central compartment at one apex and the two side arms at the other two. The cell was fitted into a 1500 ml. beaker which had an inlet at the top on one side and an outlet at the bottom of the other. The beaker with the cell was placed on a magnetic stirrer which provided the solution stirring. The rate of stirring could be varied

from 0 to 900 r.p.m. before solution cavitation occurred. The temperature was maintained at $25.0^{\circ} \pm 0.1^{\circ}\text{C}$ by means of a Wilkins Low Temperature unit used with a large capacity water bath. Water from this bath was circulated through the beaker which contained the electrolysis cell.

The recorder in the amperometric circuit was calibrated with respect to the amount of bromine generated by the constant current source. Bromine was generated in an acid solution of sodium bromide for a short period of time -- e.g., 4 sec. The recorder, at a sensitivity of 1 microamp full scale, deflected rapidly and then leveled off. More bromine was generated and the recorder response was followed. In this manner a calibration curve which consisted of a number of steps was obtained. Each step corresponded to a known generation time and was related to the molarity of bromine in solution by equations I-2 and I-3.

$$\text{Rate of generation in M/sec} = R = i/nFV \quad (\text{I-2})$$

$$RXt \text{ (sec)} = \text{total molarity generated} \quad (\text{I-3})$$

A recorder calibration graph is obtained by plotting recorder deflection in mm. vs molarity as given in equation (I-3). This plot yields a straight line. This straight line should pass through zero. However, because of solution impurities which consume bromine, the first

point is often low and the straight line passes above zero. For purposes of calculation another line is drawn parallel to the experimental line and through zero. This is done since the impurities usually react at the beginning of the generation period. Solution impurities of about $0.5 \times 10^{-6} \text{M}$ can be tolerated. This calibration graph related the total analytical concentration of bromine, $(\text{Br}_2 + \text{Br}_3^-)$, to the recorder deflection. This is evident since the calibration curve does not change with a change in the concentration of sodium bromide over the concentration range of 0.1M to 1.0M (this was the range used in these experiments and does not mean a variation is observed outside this range). The calibration was checked before each series of kinetic experiments, and the constancy of the slope was verified.

B. Reagents and Solutions

The following compounds were of reagent grade purity and were used without further purification: perchloric acid, sodium bromide, sodium nitrate, acetone, allyl alcohol, d, l-ortho-tyrosine, and 3,5-dibromotyrosine. Chromatographically pure l-tyrosine and paper chromatographically homogeneous d-tyrosine were obtained from Mann Research Laboratories. M-nitrophenol was sublimed before use.

Stock solutions of sodium bromide and sodium nitrate (2M) were made by weighing, dissolving in water, and diluting to a known volume. Stock perchloric acid solution was prepared by diluting 70% perchloric acid by one-half. An aliquot of this solution was standardized with sodium hydroxide, and a solution strength of about 5.6M was obtained. Stock solutions of compounds to be brominated were prepared in the same manner as was sodium bromide except that the weighing was carried out on an analytical balance. Dissolution of the tyrosine solutions was faster in 10^{-3} M perchloric acid than in neutral solutions. All solutions were prepared from water which was passed through a mixed cation-anion exchange resin bed. Solutions to be used in the cell were prepared by introducing aliquots of the standard solutions into a 250 ml. volumetric flask. 100 ml. of this solution was then pipetted into the cell.

The use of sodium nitrate to keep the ionic strength constant resulted in a solution which turned slightly yellow after setting for a period of several minutes and which was quite yellow after setting about 12 hours. Therefore, when sodium nitrate was used, it was added to the solution just before the solution was pipetted into the cell. Furthermore, whenever possible sodium bromide was used to keep the ionic strength constant.

C. Current-time curves

After the cell had been charged with 100 ml. of solution, the rotating platinum micro-electrode, the magnetic stirrer, and the indicator circuit were turned on. The electrode and solution were allowed to stabilize, as evidenced by a constant indicator current, the generator was actuated, and the progress of the reaction was followed. If there is no substrate present in the solution, the recorder pen deflects rapidly and the trace makes an angle of 88° with the horizontal (Curve B Figure I-1). The trace of this blank did not vary over all sodium bromide and perchloric acid concentrations used in these experiments. If substrate is present in the solution and a curve like B in Figure I-1 is obtained, the bromination reaction is too slow and cannot be studied by this method. Curve C in Figure I-1 shows the other extreme where the reaction is so fast that no deflection is obtained. This curve is like that obtained for an amperometric titration with a polarized indicator electrode, and the reaction is too rapid to be studied by this method. A typical useable current-time curve is shown by curve A. Here there is an initial rise in the bromine concentration where the rate of generation is greater than the rate of reaction. The curve then levels off, and a region where the rate of generation is just slightly greater than rate of reaction

is obtained. After this the curve rose quite sharply again. This type of curve can be used to obtain rate constants for the reaction.

D. Effect of circuit damping on current time curves

If the indicator circuit of the Sargent Model XV Polarograph is damped while the current-time curves are being recorded, circuit equilibrium is not maintained and the recorded curve lags that which would have been obtained with no damping, Figure I-2. This Figure shows, furthermore, that over the relatively flat region of the curve no error is obtained when the damping switch is in the 2 position. Therefore if one uses only the flat portions of the curve for calculations, a damping of 2 can be used to smooth out the curves. Although the curves illustrated in Figure I-2 are both smooth, in reality the undamped curve on the scale shown would consist of a line about 1/4 inch thick due to noise. In this work a damping of 2 is used and reagent concentrations are selected to fulfill the above requirement.

E. Effect of stirring rate on current-time curves

The effect the rate of stirring has on the current-time curves was investigated and will be discussed in the tyrosine results section.

F. Method of calculation

The method used in this work to analyze current-time curves for second-order rate constant data is that originally described by Fernando et. al.¹ This method is redescribed for convenience.

Under these conditions relatively large concentrations of bromide and hydrogen ions are used. Thus during a run the concentration of these ions remains constant. The total analytical concentration of bromine, $\text{Br}_2 + \text{Br}_3^-$, can be determined at any time on the curve by relating the deflection at this point to bromine concentration by means of the calibration graph mentioned earlier. The total concentration of bromine which has reacted is obtained from equation I-4.

$$[\text{Br}_2]_{\text{react}} = ([\text{Br}_2] + [\text{Br}_3^-])_{\text{electgen.}} - ([\text{Br}_2] + [\text{Br}_3^-])_{\text{unreact.}}$$

(I-4)

The concentration of unbrominated substrate in solution can then be calculated by subtracting $[\text{Br}_2]_{\text{react.}}$ from the initial substrate concentration. The rate of the reaction can be determined by use of the glass rod technique²

¹Kozak, G. and Fernando, Q., J. Phys. Chem. 67, 811 (1963)

²Hoare, J., J. Chem. Ed. 38, 570 (1961)

to measure the tangent to the current-time curve at any point on the curve. This technique utilizes two glass rods placed in parallel contact. The rods are placed over the curve in an approximately perpendicular position, and the image of the curve is observed through the rods. The rods are then rotated slightly until the image of the curve appears unbroken. A line drawn along the glass rod is perpendicular to the curve at the point directly below the point at which the two rods come into contact. The angle of the tangent θ , is then $\theta' - 90^\circ$ where θ' is the angle the perpendicular makes with respect to the horizontal. The angle θ is then related to the rate of consumption of bromine. A value of θ equal to zero means that the rate of reaction is the same as the rate of generation, and a value of θ equal to 88° means that there is no bromine consumption. Therefore a value of θ of 44° means that the rate of consumption is one-half the rate of generation. Therefore, the function $-d[\text{Br}_2]/dt$ can be found for any θ from a plot of θ vs $-d[\text{Br}_2]/dt$. All the information needed to solve the second order differential kinetic equation is thus available.

G. Brominating species and equations

With neutral bromine present in the solution, there is the possibility that in addition to molecular

bromine, bromonium ion, Br^+ , and hypo-bromous acid will also be brominating species. These reactions are governed by the following equilibria:

$$K_{\text{eq}} = 4.3 \times 10^{-9} \text{ at } 25^\circ\text{C}^1$$

$$K_{\text{eq}} = 10^{-50} \text{ at } 25^\circ\text{C}^2$$

From these equilibria one can see that the use of fairly large concentrations of bromide and hydrogen ions eliminates the need to consider brominium ion and hypobromous acid as active brominating agents.

The reaction between bromine and bromide to form tribromide ion must, however, be considered. The formation constant of this reaction, $K_{\text{Br}_3^-}$, is not known with

$$K_{\text{Br}_3^-} = \frac{[\text{Br}_3^-]}{[\text{Br}_2][\text{Br}^-]} = 17 \quad (\text{I-7})$$

certainty for these conditions, but a value of 17 seems

¹Liebhafsky, H. A., J. Amer. Chem. Soc. 56, 1500 (1934)

²Bell, R. P., and Gelles, E., J. Chem. Soc. 1951, 2734

reasonable.¹ Other studies made under similar conditions have shown that only Br_2 , and Br_3^- need be considered as active brominating species.² The bromination by bromine and tribromide ions can occur by the parallel paths given by equation I-9 where S represents the total concentration

$$\text{Rate} = k_{\text{obs}} [\text{S}] ([\text{Br}_2] + [\text{Br}_3^-]) \quad (\text{I-8})$$

$$\text{Rate} = k_1 [\text{S}] [\text{Br}_2] + k_2 [\text{S}] [\text{Br}_3^-] \quad (\text{I-9})$$

of substrate present in the solution, which in these experiments is considered to be the neutral species, $[\text{HS}]$, and the protonated species, $[\text{H}_2\text{S}^+]$. From equations I-8 and I-9 the expression given by equation I-10 is obtained.

$$k_{\text{obs}}(1 + K_{\text{Br}_3^-} [\text{Br}^-]) = k_1 + k_2 K_{\text{Br}_3^-} [\text{Br}^-] \quad (\text{I-10})$$

Therefore, from a series of runs made at constant hydrogen ion concentrations over a range of bromide ion concentrations, the constant k_1 and k_2 can be obtained from equation I-10. A plot of $k_{\text{obs}}(1 + K_{\text{Br}_3^-} [\text{Br}^-])$ vs $[\text{Br}^-]$ will give a straight line of slope $K_{\text{Br}_3^-} k_2$ and an intercept of k_1 . These rate constants are the overall rate constants

¹Griffith, R., McKlown, M., and Winn, H., Trans. Faraday Soc. 28, 101 (1932); Scaife, D. B. and Tyrrell, H. J., J. Chem. Soc. 1958, 386

²Wilson, W. J. and Soper, F. G., J. Chem. Soc. 1949, 3376; Bell, R. P. and Rawlinson, D. J., J. Chem. Soc. 1961-63

for the bromination of both the neutral and protonated forms of the substrate.

In a manner similar to the preceding, the bromination of the neutral and protonated species of the substrate can be considered as occurring by parallel reaction paths as given by equation I-12. For simplicity the individual reactions with tribromide ion will not be taken into account. As will be shown later this makes no difference.

$$\text{Rate} = k_{\text{obs}} ([\text{Br}_2] + [\text{Br}_3^-])([\text{HS}] + [\text{H}_2\text{S}^+]) \quad (\text{I-11})$$

$$\text{Rate} = k_3 [\text{Br}_2] [\text{HS}] + k_4 [\text{Br}_2] [\text{H}_2\text{S}^+] \quad (\text{I-12})$$

From equations I-11 and I-12 equation I-13 is obtained

$$k_{\text{obs}}(1 + K_{\text{Br}_3^-} [\text{Br}^-](1 + [\text{H}^+]/K_{\text{a1}})) = k_3 + k_4 [\text{H}^+]/K_{\text{a1}} \quad (\text{I-13})$$

where K_{a1} is the acid dissociation constant of the protonated substrate, H_2S^+ . Therefore a plot of the left side of equation I-13 vs $[\text{H}^+]/K_{\text{a1}}$ will yield a straight line with a slope of k_4 and an intercept of k_3 .

I-III. RESULTS AND DISCUSSION

A. Establishment of the Method

1. Bromination of Acetone

It is desirable to determine the range of second-order bromination rate constants which can be determined under these conditions. For this reason the brominations of acetone and m-nitrophenol were carried out to establish the lower range and allyl alcohol to establish an upper limit. These compounds were selected because their bromination kinetics have been studied and rate constants are available for comparison.

The current-time curves obtained for the bromination of acetone are shown in Figure I-3. With a rate of generation of 4.825 milliamps, acetone concentrations of 0.065M and 0.086M are too low to provide useful curves (Curves A and B). At an acetone concentration of 0.098M an observed rate constant, k_{obs} of $1.59 \text{ M}^{-1}\text{sec}^{-1}$ was calculated, and k_{obs} of $3.0 \text{ M}^{-1}\text{sec}^{-1}$ was found for an acetone concentration of 0.13M. The rate of bromination of acetone is complicated by a keto-enol equilibrium and by the fact that only the enol form brominates. Therefore, k_{obs} is the product of two other constants K_{E} and k_2 , where K_{E} is the keto-enol equilibrium constant for

acetone and k_2 is the second order rate constant given by equation I-14. Bell and co-workers¹ report a value of

$$\text{Rate} = k_2 [\text{enol}] ([\text{Br}_2] + [\text{Br}_3^-]) \quad (\text{I-14})$$

$5.7 \times 10^6 \text{M}^{-1}\text{sec}^{-1}$ for k_2 using a K_E value of 2.0×10^{-6} . The largest value obtained in this study is $1.5 \times 10^6 \text{M}^{-1}\text{sec}^{-1}$. This value is in reasonable agreement with $5.6 \times 10^6 \text{M}^{-1}\text{sec}^{-1}$ reported by Bell and was obtained using slightly different conditions. The rate constant k_2 , obtained in this work, increased with an increase in acetone concentration (Table I-1) which also results in a lower bromine concentration. This is caused by the complicated nature of this reaction, and a rate constant of $5.6 \times 10^7 \text{M}^{-1}\text{sec}^{-1}$ might be attainable if a lower bromine concentration were used.

Table I-1

EFFECT OF ACETONE CONCENTRATION ON THE SECOND-ORDER
BROMINATION RATE CONSTANT

Acetone Concentration	$k_2 \times 10^{-6} \text{M}^{-1}\text{sec}^{-1}$
0.098	0.80
0.117	1.02
0.125	1.25
0.130	1.29
0.134	1.50

¹Bell, R. P. and Davis, G. G., J. Chem. Soc. 1964, 902

2. Bromination of m-nitrophenol

The current-time curve obtained for the bromination of m-nitrophenol is shown in Figure I-4. For the first 60 seconds the rate of generation is much greater than rate of bromination and the current rises rapidly. After about 90 seconds, however, the rate of bromination increases to the point where its rate is greater than the rate of generation. Furthermore, calculated values of the rate constant k_{obs} defined by equation I-8 were not constant during the reaction, but they increased with time. This compound readily dibrominates¹, and the bromo substituted compound is a stronger acid than is m-nitrophenol. The presence of the monobromo substituent in solution results in an increase in the concentration of the very reactive anionic species thereby facilitating the introduction of a second bromine atom. The shape of the curve can be explained on this basis. The value of k_{obs} calculated for the monobromination reaction was $115 \text{ M}^{-1}\text{sec}^{-1}$. This value was obtained from a plot of k_{obs} vs time by extrapolation to time zero. This was done in an attempt to eliminate the effect of the second bromination. This value agrees reasonably well with the value $100 \text{ M}^{-1}\text{sec}^{-1}$ reported by Bell for k_{obs} measured under similar conditions.

¹Bell, R. P. and Spencer, T., J. Chem. Soc. 1959, 1156; Bell, R. P. and Rawlinson, D. J., J. Chem. Soc. 1961, 63

This latter value was corrected for dibromination by means of an experimentally measured value of the rate constant for the bromination of the monobromo compound.

Thus far this study has shown that it is possible to measure observed rate constants as low as $3\text{M}^{-1}\text{sec}^{-1}$ with this equipment. However, in this region a relatively large concentration of substrate must be used; and therefore other methods would be better. Good results are obtained for observed second-order rate constants of $100\text{M}^{-1}\text{sec}^{-1}$, and the method is better suited for measurement of these fast reactions.

3. Bromination of allyl alcohol

In an attempt to specify an upper limit for the rate constant, the kinetics of bromination of allyl alcohol was studied. The observed second-order rate constant for this reaction is reported to be $10^5 \text{ M}^{-1}\text{sec}^{-1}$.¹ The current time curve observed for this reaction resembles an amperometric titration curve (curve C Figure I-1). Even when a generating current of 20 milliamps was used to generate bromine in a solution containing 10^{-6}M allyl alcohol in 1M sodium bromide, no current deflection was observed before the end point. Therefore brominations with observed rate constants of $10^5\text{M}^{-1}\text{sec}^{-1}$ and greater

¹Bell, R. P. and Atkinson, J. Chem. Soc. 1963, 3200

are beyond the range of this equipment. A generating current greater than 20 milliamps was not used because of possible damage to the glass frits in the cell.

B. Kinetic study of the bromination of tyrosine

A study of the kinetics of bromination of tyrosine was undertaken in order to study a monobromination reaction. This compound is known to monobrominate in acid solutions, and forms first the 3-bromo derivative and then the 3,5-dibromotyrosine as outlined in equation I-15. As the acidity of the reaction solution is decreased, the dibromination reaction becomes more prominent¹. Therefore acid concentrations of 0.4M to 0.002M were used in this work. The dibromination product can also be oxidized by bromine

according to the scheme in equation I-16.²

¹Yagi, Y. Michel, R. and Roche, J., J. Ann Pharm. France 11, 30 (1953)

²Witkop, B., Advances in Protein Chemistry 16, 252 (1961)

A curve similar to curve A in Figure I-1 was obtained when tyrosine was brominated in acid solution. The rate constants were calculated from an analysis of the current-time curves as outlined previously. A typical set of data obtained from one run is shown in Table I-2. As can be seen the bromination of tyrosine follows second order kinetics, first-order in bromine and first-order in tyrosine. This fact was substantiated by examination of about seventy sets of data.

The effect the rate of stirring has on the rate constants was determined for a solution which contained tyrosine ($1.16 \times 10^{-4} \text{ M}$), sodium bromide (0.5M) and perchloric acid (0.10M) when the rate of stirring was 525 r.p.m. and 900 r.p.m. The stirring rate was measured with a General Radio Company Strobotac. To help avoid measuring harmonics, one end of the stirring bar was colored. As was expected the current at the higher stirring rate was slightly larger than that at the lower stirring rate (Figure I-5). However the calibration graphs for the two stirring speeds were also different (Figure I-6). The value of k_{obs} at 525 r.p.m. was $2.15 \times 10^3 \text{ M}^{-1} \text{ sec}^{-1}$ and was $2.06 \times 10^3 \text{ M}^{-1} \text{ sec}^{-1}$ at 900 r.p.m. Therefore the rate of stirring in the above region has little effect on rate constants which have an order of magnitude of $10^3 \text{ M}^{-1} \text{ sec}^{-1}$.

The second-order rate constants for the monobromination of tyrosine show a significant decrease after about 50% of the tyrosine is brominated, and are no longer constant with time. This decrease may be attributed in part to error caused by current damping. No attempt was made, therefore, to utilize the rate data beyond the 50% reaction region. Below about 20% reaction, the amount of unreacted bromine and tribromide ion present in solution is too small to give reliable measurements.

The effect bromide ion concentration has on the bromination of tyrosine is given in Table I-3, and Figure I-7 shows a plot of $k_{\text{obs}} (1 + K_{\text{Br}_3^-} \text{Br}^-)$ vs $[\text{Br}^-]$. Figure I-7 shows a straight line of zero slope from which it can be concluded that within the limitations of the experimental method, Br_3^- is not an active brominating agent.

The effect of hydrogen ion concentration on the rate constants for tyrosine was quite unusual. As Figure I-8 shows, up to a pH of about 2 hydrogen ion has little effect on the second order rate constants. Beyond this pH, however, there is a rapid increase in the rate constants with an increase in pH. Furthermore, below a pH of 2 the rate constant defined by equation I-17 shows a slight decrease with time, and above pH of 2 this rate constant increases with time. This increase suggests

there may be some dibromination occurring above a pH of 2.

$$k_{\text{obs}} (1 + K_{\text{Br}_3^-} \text{Br}^-) = k_1 [\text{S}] [\text{Br}_2] \quad (\text{I-17})$$

If dibromination occurs, then there is a possibility that oxidation will also consume bromine (equation I-16). To check this point the bromine consumption by 3,5-dibromotyrosine was investigated.

Solutions containing 3,5-dibromotyrosine ($2.4 \times 10^{-4} \text{M}$), sodium bromide (0.2M), and hydrogen ion concentrations of 0.01M, 0.0063M, 0.004M, 0.002M, and 0.001M were brominated. The solution containing 0.001M hydrogen ion consumed bromine at a fairly fast rate. The solutions with 0.004M and 0.002M used very little bromine (the slope of the curve decreased from 88° to 86.5°), and those solutions with 0.01M and 0.0063M hydrogen ion used no bromine. Therefore, oxidation of the dibromotyrosine does not contribute significantly to the rate in the regions studied.

Table I-4 contains the data for the effect of hydrogen ion on the bromination of tyrosine. Furthermore, Figure I-9 shows a plot of $k_{\text{obs}} (1 + K_{\text{Br}_3^-} [\text{Br}^-]) (1 + [\text{H}^+]/K_{\text{a}_1})$ vs $[\text{H}^+]/K_{\text{a}_1}$. This plot is a straight line; the slope of which is k_4 , the rate constant for the bromination of the protonated species, and the intercept is k_3 , the rate constant for the neutral species. The values found for k_3 and k_4 are $3 \times 10^4 \text{ M}^{-1} \text{sec}^{-1}$ and

$2.6 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$ respectively. These data then imply that the protonated species brominates with the same rate constant as that of the neutral or Zwitter ion form. The value used for K_{a1} was $10^{-2.2}$.¹

The bromination of ortho-tyrosine yielded a second-order rate constant which was about 5 times larger than that for tyrosine. Furthermore, the effect of pH on this rate constant was identical to that for tyrosine (Figure I-8). Here again below a pH of 2 the rate constant calculated from equation I-16 decreased with time during a run; however, above pH of 2, this rate constant increased. Table I-5 gives the data for the effect hydrogen ion concentration has on the rate constant. Since the K_{a1} is unknown for this compound, no plot similar to that made for tyrosine was constructed.

As one would expect, since the transition state of this reaction is assumed to be one of higher charge, an increase in rate with an increase in ionic strength was observed. The data in Table I-6 illustrates this effect in addition to showing the consistency of measurements. Although the rate constants are given for hydrogen ion concentrations of 0.36M, 0.24M, and 0.10M, as was shown, the rate constants are not affected in this region by a change in hydrogen ion concentration.

¹Albert, A., Biochem. J. 1952, 690

The second-order rate constant for the bromination of d-tyrosine was, as expected, the same as that obtained for l-tyrosine.

In phenol bromination in the para position occurs more readily than bromination in the ortho position. Consequently the introduction of a bromine atom is expected to take place more readily in the position para to the phenolic group in o-tyrosine than in the position ortho to the phenolic group in tyrosine. Since the Zwitter ion and protonated forms of both tyrosine and ortho-tyrosine were found to brominate at about the same rate, the effect of introducing a positive charge in the side-chain of the tyrosine molecule has little effect on the electron availability in the position ortho and para to the phenolic groups in tyrosine and o-tyrosine. This is substantiated by the value determined for the second-order rate constant for the monobromination of phenol ($2 \times 10^5 \text{ M}^{-1} \text{ sec}^{-1}$)¹ which is identical with that obtained for o-tyrosine. In addition the pK_1 value for tyrosine is the same as the pK_1 value for the 3,5-dibromo compound. Therefore the effects of the ring substituents are not transmitted through the side-chain of tyrosine.

¹Kozak, G. and Fernando, Q., Anal. Chim. Acta. 26, 541 (1962)

Appendix I-A

Tables I-1 - I-6

Table I-2

Kinetic Data for the Bromination of Tyrosine

Initial Concentration of tyrosine = $127.56 \times 10^{-6} \text{M}$

Initial Concentration of NaBr = 0.64M

Initial Concentration of HClO_4 = 0.34M

<u>Time</u>	$\frac{[\text{Br}_2]}{x 10^6 \text{M}}$ <u>electro-generated</u>	$([\text{Br}_2] + [\text{Br}_3^-])$ $x 10^6 \text{M}$ <u>unreacted</u>	$[\text{Br}_2]$ $x 10^7 \text{M}$ <u>unreacted</u>	$[\text{Br}_2]$ $x 10^6 \text{M}$ <u>reacted</u>	$[\text{Tyrosine}]$ $x 10^6 \text{M}$ <u>unreacted</u>	$k_{\text{obs}} (1 + K_{\text{Br}_3^-} \text{Br}^-) 10^{-3}$ <u>$\text{M}^{-1} \text{sec}^{-1}$</u>
90	22.50	0.92	0.77	21.58	105.98	2.4
120	30.00	1.00	0.84	29.00	98.56	2.4
150	37.50	1.10	0.92	36.40	91.16	2.6
180	45.00	1.16	0.97	43.84	83.72	2.4
210	52.50	1.25	1.05	51.25	76.31	2.4
240	60.00	1.35	1.13	58.65	68.91	2.4
270	67.50	1.50	1.26	66.00	61.56	2.1
300	75.00	1.67	1.40	73.33	54.23	1.9

Table I-3Effect of $[\text{Br}^-]$ on the Rate of Bromination of TyrosineIonic strength held constant at 1.0 by addition of NaNO_3

C_{Tyrosine} $\times 10^6 \text{ M}$	$[\text{Br}^-]$ M	$[\text{H}^+]$ M	k_{obs} $\text{M}^{-1} \text{sec}^{-1}$	$k_{\text{obs}}(1+K[\text{Br}^-])$
124.68	0.32	0.25	3.4×10^3	2.2×10^4
124.68	0.40	0.25	2.9×10^3	2.3×10^4
124.68	0.48	0.25	2.3×10^3	2.1×10^4
124.68	0.56	0.25	2.3×10^3	2.4×10^4
124.68	0.64	0.25	2.1×10^3	2.5×10^4
122.40	0.80	0.24	1.8×10^3	2.6×10^4

TABLE I-4

Effect of Hydrogen Ion Concentration on the Bromination of l-Tyrosine

C_{tyrosine} $\times 10^6$	$[\text{Br}^-]$	$[\text{H}^+]$	$[\text{H}^+]/K_{a1}$	$(1 + K_{\text{Br}_3^-} [\text{Br}^-])$ $\times k_{\text{obs}} 10^{-4}$	$10^{-5} (1 + K_{\text{Br}_3^-} [\text{Br}^-])$ $k_{\text{obs}} (1 + [\text{H}^+]/K_{a1})$
127.56	0.64	0.34	54.0	2.5	13.8
127.56	0.77	0.223	36.5	2.9	10.9
127.56	0.89	0.113	17.9	2.7	5.1
127.56	0.92	0.068	10.8	3.1	3.7
127.56	0.96	0.045	7.1	2.6	2.1
127.56	0.96	0.023	3.6	3.1	1.4
127.56	0.96	0.0045	0.7	2.7	.46
127.56	0.96	0.0023	0.4	5.0	.75

Table I-5
Effect of hydrogen ion concentration on the
bromination of ortho-tyrosine

C_{tyrosine}	$[\text{Br}^-]$	$[\text{H}^+]$	$(1 + K_{\text{Br}_3^-} [\text{Br}^-])$ $k_{\text{obs}} 10^{-5}$
78.69	0.72	0.113	1.4
78.69	0.76	0.068	1.5
78.69	0.80	0.023	1.5
78.69	0.80	0.0068	2.0
78.69	0.80	0.0023	4.5

Table I-6

Effect of Ionic Strength on the Bromination of l-tyrosine

C_{Tyrosine} $\times 10^6 \text{ M}$	$[\text{Br}^-]$ M	$[\text{H}^+]$ M	μ	$10^{-4} k_{\text{obs}} (1 + K_{\text{Br}_3} [\text{Br}^-])$ $\text{M}^{-1} \text{sec}^{-1}$
102.00	0.20	0.36	0.5-0.6	2.0
122.40	0.20	0.36	0.5-0.6	1.9
118.54	0.44	0.10	0.5-0.6	1.8
118.54	0.48	0.10	0.5-0.6	1.8
100.65	0.48	0.10	0.5-0.6	2.0
112.70	0.48	0.10	0.5-0.6	1.9
112.70	0.48	0.10	0.5-0.6	2.2
120.78	0.48	0.10	0.5-0.6	1.9
122.40	0.24	0.24	1.0	3.0
122.40	0.32	0.24	1.0	2.8
122.40	0.60	0.24	1.0	2.9
111.30	0.80	0.24	1.0	2.6
111.30	0.20	0.24	1.0	2.8

Appendix I-B

Figures I-1 - I-9

Figure I-1 - Typical Current-Time Curves

Figure I-2 - Damped and Undamped Current-Time Curves for the Monobromination of 1.25×10^{-4} M tyrosine in 0.32M NaBr and 0.25M HClO_4

Sensitivity: 0.004 microamp/mm

-----Damping Switch in Off Position

_____Damping Switch in Position 2

Figure I-3 - Current-Time Curves for the Bromination of Acetone in 0.1M NaBr and 0.1M HClO_4

A. 6.5×10^{-2} M Acetone

C. 9.8×10^{-2} M Acetone

B. 8.6×10^{-2} M Acetone

D. 0.13 M Acetone

Figure I-4 - Current-Time Curve for the Bromination of $1.16 \times 10^{-3} \text{M}$ m-Nitrophenol in 0.12M NaBr and 0.056M HClO_4

Figure I-5 - Effect of Rate of Stirring on Current-Time Curves

1.5×10^{-4} M tyrosine

0.50M NaBr

0.10M HClO_4

Figure I-6 - Effect of Rate of Stirring on the Calibration Graph for 0.50M NaBr and 0.10M $HClO_4$

Figure I-7 - Effect of $[Br^-]$ on the Bromination Rate Constant of Tyrosine in 0.25M $HClO_4$, $\mu = 1.0$

Figure I-8 - Effect of pH on the Bromination of tyrosine and ortho-tyrosine

Figure I-9 - Plot of $k_{obs}(1+K_{Br_3^-}[Br^-]) (1+[H^+]/K_{a1})$ versus $[H^+]/K_{a1}$ for Tyrosine

II-I. INTRODUCTION

One important application of a bromination reaction in analytical chemistry is the quantitative determination of metal ions by the dibromination of 8-quinolinol¹. In this method the chelates of 8-quinolinol are precipitated and washed free of ligand. The ligand-metal bonds are ruptured by dissolution in acid solution, and finally the 8-quinolinol is quantitatively dibrominated. This procedure for the determination of metal ions is widely used with the chelating agents 8-quinolinol and 2-methyl-8-quinolinol. The application of this method to the chelates of 4-methyl-8-quinolinol has received little attention.

In a recent paper by Graham and Corsini² it is reported that the dibromination of 4-methyl-8-quinolinol is slow and that the dibromination can be regarded as quantitative only after 20 minutes. They attributed the slowness of the reaction to steric hindrance to bromination in the 5-position by the methyl group in the 4-position. This raises the question of which bromination step is affected, the first or the second. If it is the second,

¹Hollingshead, R.G.W., Oxine and Its Derivatives, Vols. I-IV, Butterworths (London) 1954

²Graham, R. P. and Corsini, A., *Anal. Chim. Acta.* 28, 583 (1963)

then bromination must occur first in the 7-position. However, 5-bromo-4-methyl-8-quinolinol can be prepared from the bromination of the metal chelates of the 4-methyl-8-quinolinol¹; there is no known instance, except in the case of tropolone, where coordination to a metal ion has affected the position of electrophilic substitution². Therefore, on the basis of this indirect evidence, one would expect that bromination should occur first in the 5-position. Regardless of where substitution occurs, the overall rate of bromination must be slower than that of 8-quinolinol. A study of the kinetics of bromination of this compound, when compared with that of 8-quinolinol, should answer the question whether there is steric hindrance to bromination. Furthermore, if the bromination rate constants for the first and second bromination steps can be obtained and steric hindrance is present, then the question as to which bromination step is sterically hindered can be answered.

In another kinetic study of the bromination of the 4-methyl compound some interesting conclusions were reached³.

¹Prasad, R., Coffey, H. L. D., Fernando, Q., and Freiser, H., *J. Org. Chem.* 30, 1251 (1965)

²Fernando, Q., Advances in Inorganic and Radiochemistry, Vol. VII, Academic Press, 1965

³Kozak, G., Ph.D. Thesis, University of Arizona (1963)

This study showed that 4-methyl-8-quinolinol was brominated entirely by tribromide ion. This conclusion is of interest because in all other studies where tribromide ion is an active brominating agent, it is less reactive than bromine. Moreover, it is also proposed in the above study that only the protonated species of 4-methyl-8-quinolinol is brominated. This conclusion too is contrary to what one would expect.

The tribromide ion has been postulated to be an active brominating species in the second bromination step of 8-quinolinol but not in the first¹. This observation was made from kinetic data calculated from the slope of the current-time curves that are obtained toward the end of the bromination reaction. Due to the occurrence of simultaneous mono and dibromination reactions this method of analysis of kinetic data is undesirable. The availability of the 5-bromo-8-quinolinol made it possible to check this tribromide interpretation. Furthermore, if the tribromide ion is an active brominating species in the bromination of the monobromo compound, is it also the active brominating agent in the bromination of chloro, fluoro or other substituted 8-quinolinols?

¹Kozak, G., Fernando, Q., and Freiser, H., Anal. Chem. 36, 296 (1964)

The present study was therefore initiated in order to resolve the above questions.

II-II. EXPERIMENTAL

A. Apparatus

In addition to the apparatus described in Part I of this thesis, an E. A. I. Pace Analog Computer Model T. R.-10 and a Mosley Autograph X-Y recorder were used in this study.

B. Acid Dissociation Constants

The acid dissociation constants used in these calculations are as follows:

$$\text{8-quinolinol},^1 K_{a1} = 6.94 \times 10^{-6}$$

$$\text{4-methyl-8-quinolinol},^2 K_{a1} = 2.7 \times 10^{-6}$$

$$\text{5-bromo-8-quinolinol},^3 K_{a1} = 1.6 \times 10^{-4}$$

$$\text{5-chloro-8-quinolinol},^3 K_{a1} = 1.3 \times 10^{-4}$$

$$\text{5-fluoro-8-quinolinol},^3 K_{a1} = 8.9 \times 10^{-5}$$

¹Kozak, G., Fernando, Q., and Freiser, H., Anal. Chim. 36, 296 (1964)

²I.U.P.A.C. Dissociation Constants of Organic Bases in Aqueous Solutions, Perrin, D. D., Butterworths (London) 1964

³Ashizawa, F., and Freiser, H., Unpublished results

C. Reagents and solutions

Analytical reagent grade perchloric acid was used without further purification. Although Baker Analyzed Reagent Grade sodium bromide was pure enough to be used in Part I of this thesis, lot No. 26539 of this compound was found to be completely unsuitable for these experiments. However, after heating for 8-12 hours at 1100°F, the blank obtained from this reagent was normal. Solutions made from this reagent after heating contained a small amount of charcoal like substance which did not interfere with its usefulness. Analytical reagent grade sodium perchlorate is also much too impure for these measurements. Therefore, sodium perchlorate was prepared by neutralizing 4M solutions of perchloric acid and sodium hydroxide. Sodium perchlorate prepared in this manner was found to be pure enough to use.

Analytical reagent grade 8-quinolinol was sublimed under reduced pressure. The melting point of this purified compound was 72-74°C-literature value 72.5-73.5°C.¹

4-methyl-8-quinolinol was prepared from methyl-vinyl ketone and o-aminophenol by the Doebner von Miller reaction. The product was recrystallized from alcohol

¹Reagent Chemicals, Am. Chem. Soc. Specifications, Am. Chem. Soc. (Washington, D. C. 1950, p. 80)

and sublimed under reduced pressure. The purified product possessed a melting point of 137-138°C-reported value is 141°C.¹

The 5-bromo-8-quinolinol was prepared from the copper chelate of 8-quinolinol by brominating with NBS.² The monobromo was separated from the dibromo by selective pH precipitation. The 5-bromo-8-quinolinol was recrystallized from alcohol and possessed a melting point of 125°C-literature value 124°C.³

5-chloro-8-quinolinol obtained from Aldrich Chemical Company was recrystallized from alcohol. The melting point of the product was 122°C-literature 122-125°C.⁴

5-nitro-8-quinolinol was prepared from 5-nitroso-8-quinolinol, obtained from Aldrich Chemical Co., by the method of Petrow and Sturgeon.⁵ The product had a melting point of 178°C-literature value 173-178°C.

¹Phillips, J. P., Elberger, R. L., and Merritt, L. L., Jr., J. Amer. Chem. Soc. 71, 3986 (1949)

²Prasad, R., Coffey, H. L. D., Fernando, Q., and Freiser, H., J. Org. Chem. 30, 1251 (1965)

³Hollingshead, R. G. W., Oxine and Its Derivatives, Vol. III, pg. 670, Butterworths (London) 1956

⁴ibid pg. 656

⁵Petrow, V. and Sturgeon, B., J. Chem. Soc. 570 (1954)

5-fluoro-8-quinolinol was obtained from Dr. H. Gershon at the Boyce Thompson Institute, Yonkers, New York. The compound was prepared by the Schiemann reaction which proceeds by an intermediate diazonium tetrafluoroborate intermediate. The compound was purified by recrystallization from methanol. The purified product possessed a melting point of 109-110°C-literature value is 110°C.¹ The elemental analysis is as follows: carbon found 65.8% (theoretical 66.2%), hydrogen found 3.60% (theoretical 3.07%), and nitrogen found 8.58% (theoretical 8.59%). The N.M.R. Spectra of this compound in trifluoroacetic acid while different from the 5-bromo and 5-chloro compounds was entirely consistent with that expected from a 5-fluoro-8-quinolinol.

The method of solution preparation used in this study is given in Part I of this thesis. An ionic strength of 1.35 was maintained by the addition of sodium perchlorate. Solutions made with sodium perchlorate remained clear for weeks, whereas those made with sodium nitrate turned quite yellow in a few hours.

¹Hollingshead, R. G. W., Chem. and Ind. (London) 344, (1954)

II-III. Analog Computer Kinetic Study

A. Rate Equations

The application of analog computation to chemical kinetics has been described.¹ These references deal mainly with first order reactions where non-linear elements are not required. However, even for second order reactions the analog computer offers a convenient means of solving the required differential equations.

From an analysis of the data obtained from the current-time curves for 8-quinolinol and 4-methyl-8-quinolinol, one would suspect that dibromination was occurring along with monobromination. This in fact was suggested earlier for 8-quinolinol.² This dibromination can be represented by eight parallel second-order reactions.

¹Reilly, C. N., J. Chem. Ed. 39, A933 (1962)
Mathews, T., Chem. Eng. 71, 137-140 (1964, and Newton,
T. W., and Baker, F. B., U. S. Atomic Energy Comm. LA 2626
(1961)

²Kozak, G. Ph.D. Thesis, University of Arizona
(1963)

During any one run the concentration of bromide ion and of hydrogen ion remain constant; thus all the species in equilibrium remain in constant ratios. For any one run, therefore, all that can be measured is the overall rate of monobromination and of dibromination. This scheme can be represented by two parallel-consecutive second-order reactions

where A is the total concentration of unbrominated substrate, $[\text{HS}] + [\text{H}_2\text{S}^+]$, C is the total concentration of monobrominated substrate $[\text{HSBr}] + [\text{H}_2\text{SBr}^+]$, and D is that of the dibrominated substrate, $[\text{HSBr}_2] + [\text{H}_2\text{SBr}_2^+]$.

The above equations can be described by the following differential equations

$$dB/dt = R - k_{1obs}[A][B] - k_{2obs}[C][B] \quad (II-1)$$

$$-dA/dt = k_{1obs}[A][B] \quad (II-2)$$

$$-dC/dt = k_{1obs}[C][B] - k'_{1obs}[A][B] \quad (II-3)$$

where R is the rate of generation of B. In the system in question the initial concentration of B, B_0 , is zero, A is known, and C is zero. These equations can be readily solved on the analog computer, and from a knowledge of the concentration of B at any time t, the rate constants k_{1obs} and k_{2obs} can be obtained. The program used in this study is shown in Figure II-1.

B. Voltage and time scaling

This program requires the use of non-linear elements, namely the multipliers. When these elements are used, care must be taken to insure that sufficient voltage is present so that a large error will not result. One volt or greater is satisfactory. The difference in the magnitude of A and B makes voltage scaling necessary. The scaling procedure has been described very well elsewhere.¹ Because of the simplicity of this program only

¹Electronic Associates, Inc. Handbook of Analog Computation Publ. No. L. 800 0001 pp. 91-134

the basic variables A, B, and C and R were scaled. The variable B was underscaled since, with the slow reactions large concentrations of both A and B are present at the beginning of the reaction. This causes overloading at the amplifier following the multiplier since amplification by ten is required. For this reason B is multiplied by ten before the signal is fed into the multipliers. The voltage scaling used for 8-quinolinol is shown below.

Table II-1

Voltage scaling for the analog computer
kinetic study of 8-quinolinol

Variable	Max Value	Scale Factor	Scaled Variable
B	$5 \times 10^{-5} \text{M}$	$10\text{V}/5 \times 10^{-5} \text{M}$	$[2 \times 10^5 \text{B}]$
A	$1 \times 10^{-4} \text{M}$	$10\text{V}/1 \times 10^{-4} \text{M}$	$[1 \times 10^5 \text{A}]$
C	$1 \times 10^{-4} \text{M}$	$10\text{V}/1 \times 10^{-4} \text{M}$	$[1 \times 10^5 \text{C}]$
R	$2.5 \times 10^{-7} \text{M/sec}$	$10\text{V}/2.5 \times 10^{-7} \text{M/sec}$	$[4 \times 10^7 \text{R}]$

The scaled variable is now used in place of the variable in the differential equations II-1, II-2, and II-3. The scale factor is used to convert a concentration into a voltage. For example, if B is $1.05 \times 10^{-6} \text{M}$ then $(1.05 \times 10^{-6} \text{M}) (2 \times 10^5 \frac{\text{V}}{\text{M}}) = 0.210\text{V}$ is the voltage which represents this B concentration. In this manner the concentration-time curves may be converted to voltage-time curves.

The scaled equations which result from the introduction of the scaled variables into the differential equations are shown in equations II-4, II-5, and II-6.

$$\begin{aligned}
 [2 \times 10^5 B] = & -k_{1\text{obs}}/1 \times 10^5 \int_0^t [1 \times 10^5 A][2 \times 10^5 B] dt \\
 & -k_{2\text{obs}}/1 \times 10^5 \int_0^t [1 \times 10^5 C][2 \times 10^5 B] dt \\
 & + 1/200 \int_0^t [4 \times 10^7 R] dt
 \end{aligned} \tag{II-4}$$

$$[1 \times 10^5 A] = -1 \times 10^5 A_0 + k_{1\text{obs}}/2 \times 10^5 \int_0^t [1 \times 10^5 A][2 \times 10^5 B] dt \tag{II-5}$$

$$\begin{aligned}
 [1 \times 10^5 C] = & k_{1\text{obs}}/2 \times 10^5 \int_0^t [1 \times 10^5 A][2 \times 10^5 B] dt \\
 & -k_{2\text{obs}}/2 \times 10^5 \int_0^t [1 \times 10^5 C][2 \times 10^5 B] dt
 \end{aligned} \tag{II-6}$$

These equations show the rate constants as they appear in the scaled equations. As can be seen, the rate constants appear in four different constants, and four different attenuator systems must be used. Another constant $\frac{1}{200}$ appears before the rate of generation integral. The rate of generation, R, is represented by ten volts. This voltage is divided by 200 and a voltage of 0.05V is used.

The time over which data was taken during a run was usually about 400 seconds. This is much too long a time to use on the computer due to increased chances for drift over long time periods. Therefore time scaling is needed. Since increasing the gain across an integrator results only in a change in the time required to reach a final voltage and not the magnitude of this voltage, time scaling can be obtained by changing the integrator gain. In this study a time reduction of about ten was indicated; therefore the gain across all integrators was increased by a factor of ten.

C. Method of rate constant readout

The computer readout was obtained as a plot on an X-Y recorder which was synchronized to the computer. This synchronization was accomplished by feeding a reference voltage through an attenuator to the input of an integrator. The output of the integrator provided the voltage for the x scale of the computer. The integrator attenuator was adjusted to get an x deflection of 1 inch per 6 seconds. The voltage for the y scale was provided by the positive $[2 \times 10^5 B]$ output at an integrator.

The rate constants were obtained from the computer by a method of curve fitting. The concentration-time data were converted into voltage-time data by means of the scale factor for B (Table II-2) and plotted on the recorder chart.

The initial condition for A was set on the computer. The attenuators were adjusted to obtain an approximate fit to the voltage-time curve. The attenuation of the larger constant of each set was obtained and from this and the relation between them, the smaller constant was set. For example $k_{1 \text{ obs}}/1 \times 10^5$ was read and $k_{1 \text{ obs}}/2 \times 10^5$ was set. The trace was again observed and the adjustment procedure was repeated. This sequence was repeated until the best fit was obtained. This curve fitting procedure could have been facilitated if only two attenuator systems had been used and the smaller constant had been obtained from the larger by division.

The attenuation was obtained as a voltage in the following manner. A ten volt reference voltage was fed into the input of the attenuator system. The output of the system was fed to ground through a 10K ohm resistor. This provided the same load to the attenuators as did the amplifier since the amplifier grid is at virtual ground potential. The voltage between the attenuator arm and ground is read quite precisely by means of a Minneapolis-Honeywell Model 2730 potentiometer. The voltage divided by ten gives the attenuation. That this is true can be readily seen, for if 3.3V is measured, the voltage has been attenuated by 0.33. If 0.45V is measured, the attenuation is 0.045. From this constant the rate

constant can be calculated. The entire procedure for the calculation of the rate constants was repeated twice for each run.

The effect of including the second bromination reaction in the total reaction scheme, shown in Figures II-2 and II-3, is quite pronounced.

II-IV. RESULTS

A. General

One of the principal purposes of this study was to separate the rate constants for the mono and dibromination steps of 8-quinolinol and 4-methyl-8-quinolinol. That dibromination occurs is evident from the manner k_{obs} varies with time (Table II-3) for the second order bromination reactions. In this study, as in Part I of this thesis, k_{obs} is an observed constant and is defined by equation II-7.

$$k_{\text{obs}} = \frac{-d[\text{Br}_2]/dt}{[\text{S}]([\text{Br}_2] + [\text{Br}_3^-])} \quad (\text{II-7})$$

where S is the total concentration of substrate present in the neutral and protonated forms.

Dibromination of 8-quinolinol and 4-methyl-8-quinolinol has been assumed to occur by means of a parallel consecutive second-order reaction. This type of reaction path does not differentiate between acid forms of the substrate. The rate constants obtained, therefore, are apparent constants and are designated as $k_{1\text{obs}}$ and $k_{2\text{obs}}$. A good fit is obtained from the results of the computer program described above and the experimental

points for 8-quinolinol and 4-methyl-8-quinolinol as shown in Figures II-4 and II-5 respectively.

Another aspect of this study was to investigate the effect of certain electron-withdrawing substituents on the rates of bromination. The introduction of a bromine, chlorine, fluorine, or nitro group into the 5-position resulted in a compound which underwent only monobromination. The availability of the 5-bromo-8-quinolinol made it possible to compare the k_{2obs} obtained from the computer for 8-quinolinol with that for the 5-bromo compound. The monobromination of these compounds was found to be second order, i.e. first order in bromine and first order in substrate, to a good approximation. However, in some cases the k_{2obs} was found to decrease with time.

When a compound undergoes monobromination, the rate constants calculated by this method sometimes show a decrease with time during any one run. This was mentioned for the bromination of tyrosine where the rate constant k_{obs} showed a decrease with time after about 50% reaction.¹ This decrease varies with substrate concentration. For example in the bromination of 5-bromo-8-quinolinol, a slight decrease is observed when a substrate concentration of 2.0×10^{-4} M is used. However at a substrate concentration of 3.2×10^{-4} M no decrease is observed

¹O'Dom, G. and Fernando, Q., Anal. Chim. 37, 893 (1965)

over the same reaction times (60 to 360 seconds). This rate constant decrease with time is very exaggerated for the bromination of 5-nitro-8-quinolinol as is shown by the data in Table II-4. Because of the large reaction rate, a small concentration of substrate had to be used. Under these conditions the curve rises rapidly. These results show that care must be taken when evaluating data obtained by this method under conditions where there is little or no flat region in the current-time curve, since in these regions the rate constant appears to decrease with time.

As was mentioned the rate of bromination of the 5-nitro-8-quinolinol is very rapid. This appears to be opposite to that which one would expect since the nitro group is known to be strongly deactivating in aromatic electrophilic substitutions.

The strong inductive effect of the nitro group, however, causes an increase in the pK_{a1} of the 8-quinolinol. This effect is shown by the data for quinoline in Table II-5. The neutral species brominates at a much faster rate than the protonated species, 10^3 compared to 10^7 for 8-quinolinol; thus an increase in the concentration of neutral species can result in an increase in the observed rate of bromination. Moreover, phenol is a much stronger acid with the nitro substituent in the ring (acid-strengthening effect of about 10^3 for phenol). This

then introduces another term into the rate expression for the 5-nitro compound which is not present in the case of 8-quinolinol. On this basis, therefore, the observed increase in rate can be explained.

Table II-5
Acid Dissociation of Quinoline and Substituted
Quinolines in water¹ at 25°C

Substituent in Quinoline	pKa
none	4.81
5-bromo	3.62
5-chloro	3.65
5-fluoro	3.95
5-nitro	2.69

The 5-nitro-8-quinolinol brominated too rapidly to be studied under the conditions suitable to study. Therefore, only two runs were made at 0.80M sodium bromide in the presence of 0.35M or 0.45M perchloric acid. As mentioned the rate constants calculated under these conditions decreased rapidly with time. Therefore the values of k_{obs} were plotted vs time and extrapolated to zero time. Observed rate constants of $2.0 \times 10^{-4} \text{M}^{-1} \text{sec}^{-1}$ and $1.4 \times 10^{-4} \text{M}^{-1} \text{sec}^{-1}$ were obtained for 0.35M and 0.45M perchloric acid respectively.

¹I.U.P.A.C., Dissociation Constants of Organic Bases in Aqueous Solutions, Perrin, D.D., Butterworths (London) 1965

B. Analog Computer Results

With the exception of the data for 5-nitro-8-quinolinol, all the data for the monosubstituted compounds were simulated on the computer. The same quality of computer curve fit was obtained for the 5-halo-8-quinolinols as was shown for the unsubstituted 8-quinolinols in Figures II-4 and II-5.

1. 8-quinolinol

The computer results obtained for 8-quinolinol are shown in Table II-6. These data show that the agreement between the two trials is quite good. For k_{1obs} the deviation is usually $\pm 0.1 \times 10^3$ or less, and for k_{2obs} it is usually $\pm 0.3 \times 10^3$ or less. Examination of the values obtained for k_{1obs} and k_{2obs} for any one run reveals that in every case k_{2obs} is only slightly smaller than k_{1obs} . The explanation for this rather surprising result is similar to that used in the case of 5-nitro-8-quinolinol.

The overall rate constants can be analyzed for the rate constants for the bromination by bromine, k_1 , and for the bromination by tribromide ion k_2 . This analysis has been outlined in Part I of this thesis and is obtained from a plot of $k_{1obs} (1 + K_{Br_3^-} [Br^-])$ vs. $[Br^-]$. The 8-quinolinol data given in Table II-7 is plotted in Figure II-6. The plot shows a straight line of zero

slope. Therefore, at a constant $[H^+]$ of 0.23M, k_2 is zero, which means tribromide is not active in this bromination, and k_1 is $2.0 \times 10^4 M^{-1} sec^{-1}$.

With the bromide ion concentration held constant, the system can be studied to determine the rate constants for the bromination of the neutral and protonated species. The only difference between k_3 and k_4 in Part I and Part II of this thesis is that in Part I k_3 and k_4 are the constants for the bromination by molecular bromine and these in Part II are for the sum of bromine and tribromide ion. Again the reactions are considered to occur in a parallel fashion given by equation II-9.

$$\text{Rate} = k_{\text{obs}}([Br_2] + [Br_3^-])[HS] + [H_2S^+]) \quad (\text{II-8})$$

$$\text{Rate} = k_3([Br_2] + [Br_3^-])[HS] + k_4([Br_2] + [Br_3^-])[H_2S^+] \quad (\text{II-9})$$

$$k_{\text{obs}}(1 + K_{a1}/[H^+]) = k_3 K_{a1}/[H^+] + k_4 \quad (\text{II-10})$$

$$\text{When } K_{a1}/[H^+] \ll 1, \text{ then } k_{\text{obs}} = k_3 K_{a1}/[H^+] + k_4 \quad (\text{II-11})$$

A plot of k_{obs} vs $K_{a1}/[H^+]$ will give k_3 and k_4 .

A similar expression can be derived when the neutral species is kept in the expression (equation II-12)

$$k_{\text{obs}}(1 + [H^+]/K_{a1}) = k_3 + k_4[H^+]/K_{a1} \quad (\text{II-12})$$

When $1 \ll [H^+]/K_{a1}$

$$k_{\text{obs}}[H^+]/K_{a1} = k_3 + k_4[H^+]/K_{a1} \quad (\text{II-13})$$

Whenever possible both plots have been made for comparison. Furthermore, when these expressions are applied to the monohalogen compounds, all the constants are primed (i.e. k_3 is k'_3 , etc.).

Plots of k_{lobs} vs $K_{\text{al}}/[\text{H}^+]$ and $([\text{H}^+]/K_{\text{al}}) k_{\text{lobs}}$ vs $[\text{H}^+]/K_{\text{al}}$ for 8-quinolinol are shown in Figures II-7 and II-8. From these plots, values of $8 \times 10^2 \text{M}^{-1} \text{sec}^{-1}$ and $2.5 \times 10^7 \text{M}^{-1} \text{sec}^{-1}$ are obtained for the rate constants of the protonated and neutral species respectively.

2. 5-bromo-8-quinolinol

The results obtained from the bromination of 5-bromo-8-quinolinol are given in Table II-8. In this table are listed both values of the rate constants obtained from an analysis of the current-time curves and from the computer. The values obtained from the computer are slightly greater than the calculated values. The computer values have been used throughout this study. Plots of $k_{2\text{obs}} (1 + K_{\text{Br}_3} [\text{Br}^-])$ vs $[\text{Br}^-]$ are shown in Figure II-9. This figure includes the data obtained for 0.12M and 0.23M perchloric acid. The plot for 0.12M is a straight line with a positive slope of 12×10^3 which results in a rate constant, k'_2 , of $7 \times 10^2 \text{M}^{-1} \text{sec}^{-1}$. The slope of the line for 0.23M is 4×10^3 which gives a rate constant k'_2 of $2.4 \times 10^2 \text{M}^{-1} \text{sec}^{-1}$. The overall rate constants for

bromination by neutral bromine are $2.4 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$ and $1.6 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$ for acid concentrations of 0.12M and 0.23M respectively.

Values for the rate constants k_3' and k_4' of $2.0 \times 10^6 \text{ M}^{-1} \text{ sec}^{-1}$ and $1 \times 10^2 \text{ M}^{-1} \text{ sec}^{-1}$ respectively are obtained from Figure II-10. The value for k_4' is somewhat uncertain because of the small intercept. The scatter of data prevented a useful plot of $([\text{H}^+]/K_{a1})k_{2\text{obs}}$ vs $[\text{H}^+]/K_{a1}$ being made.

3. 5-chloro-8-quinolinol

The 5-chloro-8-quinolinol data, shown in Table II-9, parallel those for the 5-bromo derivative. Like the 5-bromo the 5-chloro compound also shows a slope to the plot in Figure II-11 from which the rate constant k_2' is found to be $9 \times 10^2 \text{ M}^{-1} \text{ sec}^{-1}$ and k_1' is $1.6 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$. From Figures II-12 and II-13, k_3' and k_4' are found to be $1.6 \times 10^6 \text{ M}^{-1} \text{ sec}^{-1}$ and $2.5 \times 10^2 \text{ M}^{-1} \text{ sec}^{-1}$ respectively.

4. 5-fluoro-8-quinolinol

To complete the mono-halogen series studied, the 5-fluoro data are presented in Table II-10 and Figures II-14, II-15, and II-16. A slight slope is obtained from the plot of $k_{2\text{obs}}(1+K_{\text{Br}_3^-}[\text{Br}^-])$ vs $[\text{Br}^-]$. This slope results in a rate constant for the bromination by tribromide ion of $70 \text{ M}^{-1} \text{ sec}^{-1}$. Figures II-16 and II-17, moreover,

give values of 6.5×10^5 and $1.7 \times 10^2 \text{M}^{-1} \text{sec}^{-1}$ for the rate constants k_3^1 and k_4^1 respectively.

5. 4-methyl-8-quinolinol

The rate constant values, obtained from the computer, of $k_{1\text{obs}}$ and $k_{2\text{obs}}$ for 4-methyl-8-quinolinol are given in Tables II-11 and II-14. Just as was found for 8-quinolinol, good agreement is noted between the two trials for $k_{1\text{obs}}$, and the agreement for $k_{2\text{obs}}$ is just slightly worse. Here again $k_{2\text{obs}}$ is noted to be usually slightly smaller than $k_{1\text{obs}}$. A comparison of $k_{1\text{obs}}$ from the computer and $k_{1\text{obs}}^*$ calculated from the analysis of the current-time curves shows $k_{1\text{obs}}^*$ is slightly greater. The value for $k_{1\text{obs}}^*$ was taken at the lowest time possible on the current-time curves, usually at 90 or 120 seconds. When k_{obs}^* is plotted against time and extrapolated to time equal to zero, a somewhat lower value is obtained, and this value agrees quite well with the value obtained from the computer (Table II-15). It seems possible, therefore, to obtain $k_{1\text{obs}}$ by this technique.

The data presented in Tables II-16 and II-17 have been plotted in Figures II-17, II-18, and II-19. Figure II-17 shows three plots $k_{1\text{obs}} (1 + K_{\text{Br}_3^-} [\text{Br}^-])$ vs $[\text{Br}^-]$, one for each acid concentration used. As can be seen a slightly negative slope is obtained at 0.34M and 0.23M acid and a line of zero slope for 0.12M acid. The rate

constants for the bromination by bromine are $1.6 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$, $2.0 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$ and $3.1 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$ for 0.35M, 0.23M, and 0.12M acid respectively. The rate constants k_3 and k_4 were found, from Figures II-18 and II-19, to be $7.5 \times 10^7 \text{ M}^{-1} \text{ sec}^{-1}$ and $7.5 \times 10^2 \text{ M}^{-1} \text{ sec}^{-1}$ respectively.

Since good agreement is obtained for the value of $k_{2\text{obs}}$ from the computer for 8-quinolinol and 5-bromo-8-quinolinol, the computer values for $k_{2\text{obs}}$ obtained for 4-methyl-8-quinolinol are used in calculations with confidence. The data obtained are listed in Tables II-18 and II-19. Table II-19 lists values which were calculated using K'_{a1} , the first acid dissociation constant for the monobromo-4-methyl-8-quinolinol. This value is not known, and therefore the value for K'_{a1} , the dissociation constant for 5-bromo-8-quinolinol has been used in these calculations. Figure II-20 shows plots of $k_{2\text{obs}}(1 + K_{\text{Br}_3^-} [\text{Br}^-])$ vs $[\text{Br}^-]$ for acid concentrations of 0.12M, 0.23M, and 0.35M. A positive slope is obtained in each of these cases. The slope and, therefore, the rate constant, k'_2 , increases with a decrease in hydrogen ion concentration. The rate constant k'_2 for acid concentrations of 0.12M, 0.23M, and 0.34M are respectively $1.4 \times 10^3 \text{ M}^{-1} \text{ sec}^{-1}$, $3 \times 10^2 \text{ M}^{-1} \text{ sec}^{-1}$ and $1.5 \times 10^2 \text{ M}^{-1} \text{ sec}^{-1}$. The k'_1 's for these acid concentrations are $1.0 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$, $1.6 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$, and $1.3 \times 10^4 \text{ M}^{-1} \text{ sec}^{-1}$. From Figures II-20 and II-21, values of k'_3 and k'_4 of $1.5 \times 10^6 \text{ M}^{-1} \text{ sec}^{-1}$ and $8 \times 10^2 \text{ M}^{-1} \text{ sec}^{-1}$ were found.

II-V. DISCUSSION

The results of this study have shown that the analog computer is an effective and convenient method of separating the first and second bromination rate constants of the 8-quinolinols. This is supported by the fact that the value obtained for $k_{1\text{obs}}$ from the computer agrees quite well with that obtained by extrapolation of the plot $k_{1\text{obs}}^*$, calculated from the current-time curves, vs time to time equal zero (Table II-15). Furthermore, a comparison of the rate constant for the second bromination step of 8-quinolinol with that obtained for the bromination of 5-bromo-8-quinolinol shows them to be almost the same. This lends confidence to the assumed mechanism and the method used to obtain the rate constants.

The reaction path assumed in this study does not differentiate between bromination in the 5- and 7-positions of 8-quinolinol and 4-methyl-8-quinolinol. Therefore, the rate constant obtained for the first and second brominations may be mixed constants containing both constants for the 5- and 7-positions. However, one would expect the position para to the hydroxy group to be more active in bromination reactions.¹ It will be assumed in this study

¹DeLaMare, P.B.D. and Vernon, C.A., J. Chem. Soc. 1951, 1764; Stock, L. M. and Brown, H.C., J. Amer. Chem. Soc. 82, 1942, (1960)

that the first bromine substitutes in the 5-position and the second goes in the 7-position. Therefore, when the rate constants for the bromination of the monohalogen compounds are compared to that of the unsubstituted compounds, the rates of substitution at the 7-position are being compared to those at the 5-position. Although the observed rate constants for bromination at the 5- and 7-positions appear to be about the same (i.e. $k_{1\text{obs}} \sim k_{2\text{obs}}$), this does not mean that the reactivity at the two positions is about equal since there may be different concentrations of neutral and protonated species present which brominate at different rates.

As was mentioned in the previous section, a slightly negative slope is obtained when $k_{1\text{obs}}(1 + K_{\text{Br}_3^-}[\text{Br}^-])$ is plotted vs $[\text{Br}^-]$ for 4-methyl-8-quinolinol. This behavior cannot be accounted for by the proposed reaction equations. This type of behavior was noted previously and was attributed to specific ion effects.¹ Furthermore, when sodium nitrate is used in place of sodium perchlorate to keep the ionic strength constant, a slope of zero is obtained for the above plots at a hydrogen ion concentration of 0.3M. In 0.12M acid, however, a positive slope is obtained for sodium nitrate but not for sodium

¹Bell, R. P. and Rawlinson, D. J., J. Chem. Soc. 1961, 63

perchlorate. It seems evident, therefore, that specific ion effects are present in these reactions. The results show that straight line plots with positive slopes are obtained when $k_{2\text{obs}} (1 + K_{\text{Br}_3^-} [\text{Br}^-])$ vs $[\text{Br}^-]$ is plotted for 4-methyl-8-quinolinol (Figure II-20). These slopes, furthermore, increase as the hydrogen ion concentration is decreased. This is not observed for the monobromination of 4-methyl-8-quinolinol where the above plots give a slope of zero. Moreover, tribromide is not active in the monobromination of 8-quinolinol at a hydrogen ion concentration of 0.23M. However, here again tribromide ion is active in the bromination of 5-bromo-8-quinolinol, and this activity is greater at 0.12M hydrogen ion than at 0.23M hydrogen ion. Furthermore, a positive slope is found for the plot of $k_{2\text{obs}} (1 + K_{\text{Br}_3^-} [\text{Br}^-])$ vs $[\text{Br}^-]$ in the bromination of the 5-chloro compound at 0.12M acid and the 5-fluoro compound at 0.23M acid. Thus it appears that in every case tribromide is active in the bromination of the monohalo-8-quinolinols but not in the monobromination of 8-quinolinol.

In these experiments the tribromide-ion activity is observed for the slower halogen substituted-8-quinolinols reactions but not for the faster unsubstituted 8-quinolinols reactions. This is contrary to that normally observed. For example, no tribromide activity is found for the bromination of neutral m-nitrophenol

($k = 100\text{M}^{-1}\text{sec}^{-1}$). However, the phenoxide anion shows a tribromide ion rate constant of $1.3 \times 10^8 \text{M}^{-1}\text{sec}^{-1}$ and a bromine rate constant of $1.3 \times 10^9 \text{M}^{-1}\text{sec}^{-1}$.¹ Furthermore, tribromide is active in the bromination of the enol form of acetone, $k_2 \sim 10^7 \text{M}^{-1}\text{sec}^{-1}$,² and the enols of diethyl malonate and methyl methantricarboxylate³ all of which have rate constants $10^6 \text{M}^{-1}\text{sec}^{-1}$ or greater. However, no tribromide ion activity is observed in the bromination of anisoles which are fairly slow, $k < 10^4 \text{M}^{-1}\text{sec}^{-1}$.¹

The effect that an electron withdrawing substituent has on the rate of bromination at the position meta to the substituent was investigated for three halogens. The general effect, as expected, is to deactivate the phenolic ring. The 5-bromo-8-quinolinol is found to brominate at a rate which was approximately ten times slower than the 8-quinolinol. As was pointed out, however, this is a comparison of activities at two different positions. The bromination of the 5-chloro compound is approximately the same as that of the 5-bromo ($k_3 = 1.8 \times 10^6 \text{M}^{-1}\text{sec}^{-1}$ compared

¹Bell, R. P. and Spencer, T., J. Chem. Soc. 1959, 1156

²Bell, R. P. and Davis, G. G., J. Chem. Soc. 1964, 902

³Bell, R. P. and Rawlinson, D. J., J. Chem. Soc. 1961, 726

to $k'_3 = 2.0 \times 10^6 \text{ M}^{-1} \text{ sec}^{-1}$ for the 5-bromo compound). The 5-fluoro-8-quinolinol, however, brominates at a rate which is about three times slower than that of the 5-bromo-8-quinolinol ($k'_3 = 7.3 \times 10^5 \text{ M}^{-1} \text{ sec}^{-1}$ compared to $k'_3 = 2.0 \times 10^6 \text{ M}^{-1} \text{ sec}^{-1}$). This is unexpected in view of published data for phenols and naphthalenes.¹ In that study it was found that for the bromination of phenyl halides, the rate constant for the bromination of the fluoride substituent was about the same as the hydrogen substituent, and the bromide and chloride were very deactivating. This of course was for compounds whose bromination occurs ortho-para to the halogen. The reactivities of the halogen substituted 8-quinolinols may be examined qualitatively by using the Brown sigma-plus meta values, σ_m^+ , obtained for the benzene ring. The values for fluoro, chloro, and bromo respectively are 0.337, 0.373, and 0.391. The σ_m^+ is defined in equation II-14, where k is the rate constant for the substituent in question, k_o is the rate constant of the reference and ρ is a constant which depends on the compound in question. By use of this equation it can be shown that the rate constant for the chloro substituted compound should be about the same as that for the bromo

$$\log k/k_o = \rho \sigma_m^+ \quad (\text{II-14})$$

¹DeLaMare, P.B.D. and Robertson, P.W., J. Chem. Soc. 1948, 100

which should be less than that for the fluoro substituted compound. This is not what was observed in this study for the rate constants for the bromination of the neutral molecules. However, the data are not good or extensive enough to make any quantitative explanation. The rate constants for the protonated species of the three compounds are about the same.

The dibromination of 4-methyl-8-quinolinol yielded results which differ greatly from those of two previous studies. In one study¹ it was found that 4-methyl-8-quinolinol was brominated entirely by the tribromide ion. Furthermore, it was found that only the protonated species brominated. These facts were in sharp contrast to the results found for the bromination of 8-quinolinol and 2-methyl-8-quinolinol. In the other study² it was found that dibromination of 4-methyl-8-quinolinol was relatively slow compared to that of 8-quinolinol.

This study gave quite different results for the bromination of 4-methyl-8-quinolinol. It was found that there is little difference between the behavior of this compound and that of 8-quinolinol. Just as is observed for 8-quinolinol, the observed second bromination constant,

¹Kozak, G., Ph.D. Thesis University of Arizona
(1963)

²Corsini, A. and Graham, R. P., Anal. Chim. Acta.
28, 583 (1963)

$k_{2\text{obs}}$, is approximately the same magnitude as $k_{1\text{obs}}$. Furthermore, these constants are somewhat larger than those found for 8-quinolinol. The bromination of 4-methyl-8-quinolinol, moreover, is very dependent on the hydrogen ion concentration in the range studied (from 0.05 to 0.47M), and the neutral species brominates at a much greater rate than does the protonated species. The bromination rate constants of 4-methyl-8-quinolinol are compared to those of 8-quinolinol in Table II-20. The rate constants are defined by the reaction scheme given below for 4-methyl-8-quinolinol.

As can be seen from Table II-20, the rate constants for the monobromination of 4-methyl-8-quinolinol are even larger than those for 8-quinolinol.

Table II-20

Comparison of the monobromination rate constants
of 4-methyl-8-quinolinol and 8-quinolinol

Rate constants $M^{-1}sec^{-1}$	8-quinolinol	4-methyl- 8-quinolinol
k_1	2.0×10^4	2.0×10^4
k_2	0	0
k_3	2.6×10^7	7.5×10^7
k_4	8×10^2	7.2×10^2

The results for the second bromination step of 4-methyl-8-quinolinol are compared to those of 5-bromo-8-quinolinol in Table II-21. These values are about equal, within experimental error. The calculations of k'_3 and k'_4 for the 4-methyl compound were made assuming k'_{a1} to be the same as that for 5-bromo-8-quinolinol, and these constants may be in error.

Table II-21

Comparison of the second bromination rate constants
of 4-methyl-8-quinolinol and 5-bromo-8-quinolinol

Rate constants $M^{-1}sec^{-1}$	8-quinolinol	4-methyl 8-quinolinol
k'_1	1.95×10^4	1.6×10^4

Table II-21 - Continued

Rate constants $M^{-1}sec^{-1}$	8-quinolinol	4-methyl 8-quinolinol
k'_2	0	3×10^2
k'_3	2.0×10^6	1.6×10^6
k'_4	-	7.5×10^2

There is no simple explanation for the observed difference between this work and the Corsini and Graham study except that they were made under different conditions. The Corsini and Graham work was carried out in aqueous acetic acid by classical methods.

The observed difference between this study and the Kozak study cannot be attributed to conditions, for they were both made under about the same conditions. One difference between the two is that although approximately the same generation rate was used in the two studies, the solution volume used in the Kozak experiments was one-half that used in this study. Therefore the actual rate of generation of the previous study was twice that of this study. This could result in different shaped current-time curves, and as was described earlier, curves which rise rapidly with time yield rate constants which decrease with time. This type of curve was used by Kozak. It seems possible, therefore, that the difference between the two experimental results might be attributed to this.

On the basis of the above results there is no evidence for steric hindrance to bromination. The bromination of 4-methyl-8-quinolinol occurs with rate constants which are comparable to that of 8-quinolinol. Furthermore, tribromide is not active in the monobromination of 4-methyl-8-quinolinol, and only slightly active in the dibromination step.

Appendix II-A
Tables II-2 - II-19

TABLE II-2

Concentration-time and voltage-time data for 8-quinolinol

Initial concentration of 8-quinolinol $1.00 \times 10^{-4} \text{M}$

Initial concentration of NaBr = 0.64M

Initial concentration of $\text{HClO}_4 = 0.23 \text{M}$

Time	$10^6 ([\text{Br}_2] + [\text{Br}_3^-])$	$[2 \times 10^5 \text{B}]$
90	1.57	0.314V
120	1.57	0.314V
180	1.61	0.322V
240	1.72	0.344V
300	1.88	0.376V
360	2.10	0.420V
420	2.35	0.470V

TABLE II-3

Example of the Variation of k_{obs}^* with Time of $1.36 \times 10^{-4} \text{M}$
4-methyl-8-quinolinol in 0.64M NaBr

Initial Concentration of $\text{HClO}_4 = 0.05\text{M}$

Time	$10^{-3} k_{\text{obs}}^*$	Time	$10^{-3} k_{\text{obs}}^*$
90	4.8	240	7.0
120	5.1	270	7.6
150	5.5	300	8.3
180	6.0	330	9.2
210	6.5	360	10.1

Initial Concentration of $\text{HClO}_4 = 0.35\text{M}$

Time	$10^{-3} k_{\text{obs}}^*$	Time	$10^{-3} k_{\text{obs}}^*$
120	1.6	270	2.2
150	1.7	300	2.4
180	1.8	330	2.5
210	1.9	360	2.8
240	2.1		

TABLE II-4

Variation of k_{obs} with time for 5-nitro-8-quinolinol

0.35M HClO_4 0.80M NaBr		0.45M HClO_4 0.80M NaBr	
t (sec)	$10^{-3}k_{\text{obs}}^*$	t (sec)	$10^{-3}k_{\text{obs}}^*$
60	8.14	60	5.72
90	7.25	90	4.93
120	6.25	120	4.13
150	4.81	150	3.38
180	3.36	180	2.40

TABLE II-6

Analog Computer Results for 1.00×10^{-4} M 8-quinolinolInitial Concentration of $\text{HClO}_4 = 0.23\text{M}$

[NaBr]		Trial		Average
		1	2	
0.24	$10^{-3}k_{1\text{obs}}$	3.48	3.50	3.49
	$10^{-3}k_{2\text{obs}}$	3.38	3.28	3.33
0.64	$10^{-3}k_{1\text{obs}}$	1.65	1.85	1.75
	$10^{-3}k_{2\text{obs}}$	1.55	1.31	1.43
0.80	$10^{-3}k_{1\text{obs}}$	1.34	1.40	1.37
	$10^{-3}k_{2\text{obs}}$	1.30	1.21	1.25
1.00	$10^{-3}k_{1\text{obs}}$	1.16	1.15	1.16
	$10^{-3}k_{2\text{obs}}$	0.94	0.96	0.95

Initial Concentration of NaBr 0.64M

[HClO ₄]		1	2	Average
0.05	$10^{-3}k_{1\text{obs}}$	4.60	4.60	4.60
	$10^{-3}k_{2\text{obs}}$	3.30	3.00	3.15
0.12	$10^{-3}k_{1\text{obs}}$	2.60	2.60	2.60
	$10^{-3}k_{2\text{obs}}$	1.84	2.04	1.94
0.23	$10^{-3}k_{1\text{obs}}$	1.65	1.85	1.75
	$10^{-3}k_{2\text{obs}}$	1.55	1.31	1.43

TABLE II-6 - Continued

[HClO ₄]		1	2	Average
0.35	$10^{-3}k_{1\text{obs}}$	1.34	1.35	1.35
	$10^{-3}k_{2\text{obs}}$	1.30	1.31	1.31

TABLE II-7

The Effect of Bromide and Hydrogen Ions on the Bromination of $1.00 \times 10^{-4} M$
8-quinolinol

[HClO ₄]	[NaBr]	$10^{-3} k_{2obs}$		$10^{-3} k_{2obs}$ ($1 + K_{Br_3} + [Br]$)	$10^{-4} [H^+] / K_{a1}$	$10^5 K_{a1} / [H^+]$
		Calc.	Computer			
0.23	0.24	3.91	3.49	17.7	--	--
0.23	0.64	1.97	1.75	20.8	--	--
0.23	0.80	1.60	1.37	20.0	--	--
0.23	1.00	1.97	1.16	20.9	--	--
0.05	0.64	5.48	4.60	--	0.72	13.88
0.12	0.64	3.02	2.60	--	1.73	5.78
0.23	0.64	1.97	1.75	--	3.31	3.02
0.35	0.64	1.74	1.35	--	5.04	1.98

TABLE II-8

The Effect of Bromide and Hydrogen Ions on the Bromination of $2.01 \times 10^{-4} \text{M}$

5-bromo-8-quinolinol

[HClO ₄]	[NaBr]	$10^{-3}k_{2\text{obs}}$		$10^{-3}k_{2\text{obs}}$ ($1 + K_{\text{Br}_3} - [\text{Br}]$)	$10^{-2}[\text{H}^+]/K_{\text{a1}}$	$10^4(K_{\text{a1}}/[\text{H}^+])$
		Calc.	Computer			
0.12	0.24	5.28	5.78	26.8	--	--
0.12	0.64	2.29	2.58	30.7	--	--
0.12	0.80	2.05	2.33	34.0	--	--
0.12	1.00	1.78	2.01	36.2	--	--
0.23 ^a	0.24	3.30	3.27	16.6	--	--
0.23 ^a	0.64	1.51	1.50	17.8	--	--
0.23 ^a	0.80	1.19	1.28	18.6	--	--
0.23 ^a	1.00	0.98	1.07	19.3	--	--
0.05	0.64	5.84	6.30	--	4.3	32.0
0.12	0.64	2.29	2.58	--	10.3	13.3
0.23	0.64	1.51	1.28	--	19.8	6.95
0.35	0.64	0.84	1.10	--	30.2	4.57

 $a = 2.41 \times 10^{-4} \text{M}$ 5-bromo-8-quinolinol

TABLE II-9

The Effect of Bromide and Hydrogen Ions on the Bromination of 1.60x

 10^{-4} M 5-chloro-8-quinolinol

[HClO ₄]	[NaBr]	$10^{-3}k_{2obs}$		$10^{-3}k_{2obs}$ ($1+K_{Br3^-}[Br]$)	$10^{-2}[H^+]/K_{al}$	$10^4K_{al}/[H^+]$
		Calc.	Computer			
0.12	0.12	5.99	5.74	17.4	--	--
0.12	0.24	3.82	3.84	19.5	--	--
0.12	0.32	3.29	3.22	20.7	--	--
0.12	0.64	1.91	2.16	25.7	--	--
0.12	0.80	1.62	1.87	27.3	--	--
0.12	1.00	1.24	1.42	25.6	--	--
0.05	0.64	4.19	4.02	--	3.8	26.0
0.12	0.64	1.91	2.16	--	9.2	10.8
0.23	0.64	0.90	1.13	--	17.7	5.65
0.35	0.64	0.58	0.82	--	26.9	3.72

TABLE II-10

The Effect of Bromide and Hydrogen Ions on the Bromination of 5-fluoro
8-quinolinol

[HClO ₄]	[NaBr]	10 ⁻³ k _{2obs}		10 ⁻³ k _{2obs} (1+K _{Br₃⁻ [Br]}	10 ⁻² [H ⁺]	10 ⁴ K _{a1}
		Calc.	Computer		$\frac{K'_{a1}}{[H^+]}$	$\frac{K_{a1}}{[H^+]}$
0.23 ^a	0.24	0.76	0.80	4.06	--	--
0.23 ^a	0.32	0.60	0.66	4.25	--	--
0.23 ^a	0.64	0.30	0.36	4.28	--	--
0.23 ^b	0.80	0.30	0.34	4.96	--	--
0.23 ^b	1.00	0.23	0.28	5.04	--	--
0.05 ^c	0.64	1.28	1.26	--	5.6	17.8
0.12 ^c	0.64	0.62	0.66	--	13.4	7.41
0.17 ^c	0.64	0.48	0.53	--	19.1	5.23
0.23 ^c	0.64	0.36	0.43	--	25.8	3.87
0.35 ^c	0.64	0.27	0.32	--	39.3	2.54

[5-fluoro-8-quinolinol], a = 2.49x10⁻⁴M; b = 3.74x10⁻⁴M; c = 4.15x10⁻⁴M

TABLE II-11

Analog Computer Results for 4-methyl-8-quinolinol
 Initial Concentration of 4-methyl-8-quinolinol $1.36 \times 10^{-4} \text{M}$

		0.12M HClO ₄		
		Trial		
[NaBr]		1	2	Average
0.12	10 ⁻³ k _{1obs}	7.30	7.30	7.30
	10 ⁻³ k _{2obs}	4.28	4.27	4.28
0.24	10 ⁻³ k _{1obs}	5.84	5.95	5.89
	10 ⁻³ k _{2obs}	3.23	2.95	3.09
0.64	10 ⁻³ k _{1obs}	2.72	2.67	2.74
	10 ⁻³ k _{2obs}	2.20	2.38	2.29
0.80	10 ⁻³ k _{1obs}	2.12	2.08	2.10
	10 ⁻³ k _{2obs}	1.98	2.11	2.04
1.00	10 ⁻³ k _{1obs}	1.62	1.62	1.62
	10 ⁻³ k _{2obs}	2.01	1.79	1.90

TABLE II-12

Analog Computer Results for 4-methyl-8-quinolinol
in 0.23M HClO_4

Initial concentration of 4-methyl-8-quinolinol $1.36 \times 10^{-4}\text{M}$

		<u>Trial</u>		Average
		1	2	
0.12	$10^{-3}k_{1\text{obs}}$	6.22	6.50	6.36
	$10^{-3}k_{2\text{obs}}$	4.00	3.85	3.93
0.24	$10^{-3}k_{1\text{obs}}$	4.17	4.06	4.11
	$10^{-3}k_{2\text{obs}}$	3.41	3.41	3.41
0.64	$10^{-3}k_{1\text{obs}}$	1.72	1.58	1.65
	$10^{-3}k_{2\text{obs}}$	1.34	1.85	1.60
0.80	$10^{-3}k_{1\text{obs}}$	1.33	1.30	1.32
	$10^{-3}k_{2\text{obs}}$	1.38	1.36	1.37
1.00	$10^{-3}k_{1\text{obs}}$	1.03	1.00	1.02
	$10^{-3}k_{2\text{obs}}$	1.12	1.25	1.18

TABLE II-13

Analog Computer Results for 4-methyl-8-quinolinol
in 0.35M HClO_4

Initial concentration of 4-methyl-8-quinolinol		$1.36 \times 10^{-4} \text{M}$		
[NaBr]		Trial		Average
		1	2	
0.12	$k_{1\text{obs}} 10^{-3}$	5.15	5.19	5.17
	$k_{2\text{obs}} 10^{-3}$	4.00	3.90	3.95
0.24	$k_{1\text{obs}} 10^{-3}$	3.46	3.36	3.41
	$k_{2\text{obs}} 10^{-3}$	2.81	2.91	2.86
0.64	$k_{1\text{obs}} 10^{-3}$	1.33	1.37	1.35
	$k_{2\text{obs}} 10^{-3}$	1.38	1.18	1.28
0.80	$k_{1\text{obs}} 10^{-3}$	1.03	0.99	1.01
	$k_{2\text{obs}} 10^{-3}$	1.12	1.12	1.12
1.00	$k_{1\text{obs}} 10^{-3}$	0.76	0.79	0.78
	$k_{2\text{obs}} 10^{-3}$	0.80	0.87	0.84

TABLE II-14

Analog Computer Results for 4-methyl-8-quinolinol

in 0.64M NaBr

Initial Concentration of 4-methyl-8-quinolinol $1.36 \times 10^{-4} \text{M}$

[HClO ₄]		Trial		Average
		1	2	
0.05	$10^{-3}k_{1\text{obs}}$	4.09	4.01	4.05
	$10^{-3}k_{2\text{obs}}$	4.10	4.43	4.26
0.07	$10^{-3}k_{1\text{obs}}$	3.63	3.51	3.57
	$10^{-3}k_{2\text{obs}}$	3.20	3.50	3.35
0.095	$10^{-3}k_{1\text{obs}}$	2.90	2.95	2.92
	$10^{-3}k_{2\text{obs}}$	3.17	2.82	2.99
0.12	$10^{-3}k_{1\text{obs}}$	2.72	2.67	2.69
	$10^{-3}k_{2\text{obs}}$	2.20	2.38	2.29
0.23	$10^{-3}k_{1\text{obs}}$	1.72	1.58	1.65
	$10^{-3}k_{2\text{obs}}$	1.34	1.85	1.59
0.35	$10^{-3}k_{1\text{obs}}$	1.33	1.37	1.35
	$10^{-3}k_{2\text{obs}}$	1.38	1.18	1.28
0.47	$10^{-3}k_{1\text{obs}}$	1.11	1.12	1.12
	$10^{-3}k_{2\text{obs}}$	1.09	0.99	1.04

TABLE II-15

Comparison of k_{lobs} from the Analog Computer with k_{lobs}^* from Extrapolation to $t=0$ for $1.36 \times 10^{-4} \text{M}$

4-methyl-8-quinolinol

$[\text{HClO}_4]$	$[\text{NaBr}]$	$10^{-3}k_{\text{lobs}}$	$10^{-3}k_{\text{lobs}}^*$
0.05	0.64	4.2	4.1
0.05	0.80	3.8	3.7
0.05	1.00	2.2	2.2
0.12	0.12	7.30	7.4
0.12	0.24	5.89	5.8
0.12	0.64	2.74	2.4
0.12	0.80	2.10	1.9
0.12	1.00	1.62	1.5
0.23	0.12	6.36	6.5
0.23	0.24	4.11	4.4
0.23	0.64	1.65	1.5
0.23	0.80	1.32	1.2
0.23	1.00	1.02	0.9
0.35	0.12	5.17	5.3
0.35	0.24	3.41	3.9
0.35	0.64	1.35	1.2
0.35	0.80	1.01	1.0
0.35	1.00	0.78	0.7

TABLE II-16

The Effect of Bromide Ion Concentration on the
Bromination of $1.36 \times 10^{-4} \text{ M}$ 4-methyl-8-quinolinol

[HClO ₄]	[NaBr]	$10^{-3} k_{\text{lobs}}$		$10^{-4} k_{\text{lobs}}$ ($1 + K_{\text{Br}_3^-} [\text{Br}]$)
		Calc.	Computer	
0.12	0.12	7.88	7.30	2.22
0.12	0.24	6.59	5.89	2.99
0.12	0.64	3.00	2.74	3.26
0.12	0.80	2.54	2.10	3.07
0.12	1.00	2.03	1.62	2.92
0.23	0.12	7.00	6.36	1.93
0.23	0.24	5.00	4.11	2.09
0.23	0.64	2.05	1.65	1.96
0.23	0.80	1.61	1.32	1.93
0.23	1.00	1.24	1.02	1.84
0.35	0.12	6.33	5.17	1.57
0.35	0.24	4.10	3.41	1.73
0.35	0.64	1.61	1.35	1.60
0.35	0.80	1.23	1.01	1.48
0.35	1.00	1.00	0.78	1.40

TABLE II-17

The Effect of Hydrogen Ion Concentration on the
 Bromination of 1.36×10^{-4} M 4-methyl-8-quinolinol in 0.64M NaBr

[HClO ₄]	$10^{-3}k_{\text{obs}}$		$10^6 K_{a1}/[H^+]$	$10^{-4}[H^+]/K_{a1}$
	Calc.	Computer		
0.05	4.77	4.05	54.00	1.85
0.07	4.02	3.57	38.57	2.59
0.095	3.25	2.92	28.46	3.52
0.12	3.00	2.69	22.50	4.44
0.23	2.05	1.65	11.74	8.52
0.35	1.61	1.35	7.71	12.96
0.47	1.36	1.12	5.74	17.40

TABLE II-18

The Effect of Bromide Ion Concentration on the Second
Bromination Step of 1.36×10^{-4} M 4-methyl-8-quinolinol

$[\text{HClO}_4]$	$[\text{NaBr}]$	$10^{-3}k_{2\text{obs}}$	$10^{-3}k_{2\text{obs}}(1+K_{\text{Br}_3}-[\text{Br}^-])$
0.12	0.12	4.28	13.0
0.12	0.24	3.09	15.7
0.12	0.64	2.29	27.2
0.12	0.80	2.04	29.8
0.12	1.00	1.90	34.0
0.23	0.12	3.93	11.9
0.23	0.24	3.41	17.3
0.23	0.64	1.60	19.0
0.23	0.80	1.37	20.0
0.23	1.00	1.18	21.2
0.35	0.12	3.95	12.0
0.35	0.24	2.86	14.5
0.35	0.64	1.28	15.2
0.35	0.80	1.12	16.3
0.35	1.00	0.84	15.1

TABLE II-19

The Effect of Hydrogen Ion Concentration on the Second
 Bromination of 1.36×10^{-4} M 4-methyl-8-quinolinol in 0.64M NaBr

$[\text{HClO}_4]$	$10^{-3}k_{2\text{obs}}$	$10^{-4}k'_{a1}/[\text{H}^+]$	$10^{-2}[\text{H}^+]/K'_{a1}$
0.05	4.26	23.2	4.31
0.07	3.35	16.57	6.02
0.095	2.99	12.21	8.19
0.12	2.29	9.67	10.34
0.23	1.59	5.04	19.83
0.35	1.28	3.31	30.17
0.47	1.04	2.47	40.52

Appendix II-B
Figures II-1 - II-22

Figure II-1 - Computer Program for Consecutive Parallel Second-Order Reactions

Figure II-2 - The Effect k_2 has on the Shape of the Voltage-Time Curves for $k_1 = 1.1 \times 10^3 \text{ M}^{-1} \text{ sec}^{-1}$

Figure II-3 - The Effect k_2 has on the Shape of the Voltage-Time Curves for $k_1 = 6.2 \times 10^3 \text{ M}^{-1} \text{ sec}^{-1}$

Figure II-4 - Example of Computer Fit for 8-quinolinol in 0.64M NaBr

Figure II-5 - Example of Computer Fit for 4-methyl-8-quinolinol in 0.12M HClO₄

Figure II-6 - Plot of $k_{1\text{obs}} (1 + K_{\text{Br}_3^-} [\text{Br}^-])$ versus $[\text{Br}^-]$ for 8-quinolinol

Figure II-7 - Plot of k_{1obs} versus $K_{a1}/[H^+]$ for 8-quinolinol

Figure II-8 - Plot of $k_{\text{lobs}}([H^+]/K_{\text{al}})$ versus $[H^+]/K_{\text{al}}$ for 8-quinolinol

Figure II-9 - Plots of $k_{2\text{obs}}(1 + K_{\text{Br}_3^-} [\text{Br}^-])$ versus $[\text{Br}^-]$ for 5-bromo-8-quinolinol

Figure II-10 - Plot of k_{2obs} versus $k'_{al}/[H^+]$ for 5-bromo-8-quinolinol

Figure II-11 - Plot of $k_{2\text{obs}}(1 + K_{\text{Br}_3^-} [\text{Br}^-])$ versus $[\text{Br}^-]$ for 5-chloro-8-quinolinol in 0.12M HClO_4

Figure II-12 - Plot of k_{2obs} versus $K'_{a1}/[H^+]$ for 5-chloro-8-quinolinol

Figure II-13 - Plot of $k_{2\text{obs}} ([\text{H}^+]/K'_{a1})$ vs $[\text{H}^+]/K'_{a1}$ for 5-chloro-8-quinolinol

Figure II-14 - Plot of $k_{2obs} (1 + K_{Br_3^-} [Br^-])$ versus $[Br^-]$ for 5-fluoro-8-quinolinol

Figure II-15 - Plot of k_{2obs} versus $K'_{a1}/[H^+]$ for 5-fluoro-8-quinolinol

Figure II-16 - Plot of $k_{2\text{obs}}([H^+]/K'_{a1})$ versus $[H^+]/K'_{a1}$ for 5-fluoro-8-quinolinol

Figure II-17 - Plots of $k_{\text{obs}}(1+K_{\text{Br}_3^-} [\text{Br}^-])$ versus $[\text{Br}^-]$ for 4-methyl-8-quinolinol

Figure II-18 - Plot of k_{1obs} versus $K_{a1}/[H^+]$ for 4-methyl-8-quinolinol

Figure II-19 - Plot of $k_{1obs} [H^+]/K_{a1}$ versus $[H^+]/K_{a1}$ for 4-methyl-8-quinolinol

Figure II-20 - Plots of $k_{2\text{obs}}(1 + K_{\text{Br}_3^-} [\text{Br}^-])$ versus $[\text{Br}^-]$ for 4-methyl-8-quinolinol

Figure II-21 - Plot of k_{2obs} versus $K'_{a1}/[H^+]$ for the second bromination of 4-methyl-8-quinolinol

Figure II-22 - Plot of $k_{2\text{obs}}([H^+]/K'_{a1})$ versus $[H^+]/K'_{a1}$ for 4-methyl-8-quinolinol