2-SUBSTITUTED 8-QUINOLINOLS

AS

TERDENTATE CHELATING AGENTS

2-SUBSTITUTED 8-QUINOLINOLS
AS
TERDENTATE CHELATING AGENTS

By

RICHARD M. CASSIDY, B.Sc., M.Sc.

A Thesis
Submitted to the Faculty of Graduate Studies
in Partial Fulfilment of the Requirements
for the Degree
Doctor of Philosophy

McMaster University
September, 1971

DOCTOR OF PHILOSOPHY (1971)
(Chemistry)

McMASTER UNIVERSITY
Hamilton, Ontario

TITLE: 2-Substituted 8-Quinolinols as Terdentate Chelating Agents

AUTHOR: Richard Cassidy, B.Sc., M.Sc. (Dalhousie University)

SUPERVISOR: Dr. A. Corsini

NUMBER OF PAGES: ix, 168

SCOPE AND CONTENTS:

A series of 2-substituted 8-hydroxyquinolines have been prepared and their reactions with a number of metal ions studied. Acid-dissociation constants and metal-chelate formation constants have been determined. 2-(2'-Hydroxyphenyl), 2-(2'-pyridyl) and 2-hydrazino-8-hydroxyquinoline function as terdentate ligands. The latter two ligands form a number of protonated complexes some of which are unusually stable. This stability is likely due to hydrogen bonding.

ACKNOWLEDGEMENTS

The author wishes to thank Dr. A. Corsini for his help and advice throughout the work.

He also wishes to thank Dr. Hileman, Dr. Warkentin, Dr. Olsen, Dr. Datars and Michael Thompson for their helpful discussion.

The author also acknowledges the financial assistance from the Department of Chemistry and the Province of Ontario.

TABLE OF CONTENTS

LIST OF TABLES

APPENDIX II

LIST OF FIGURES

Number Title Page

APPENDIX IV

GENERAL INTRODUCTION

Since 1920, when Werner first proposed a ring structure for the ethylenediamine complex of Pt(II), the study of metal chelates has increased rapidly and their importance in several fields of science is now recognized(1-3). The properties of many chelates exhibit a high degree of specificity which is not fully understood. Chlorophyll and haemoglobulin are examples of the importance and specificity of certain chelates.

In chemistry, the interest in chelating agents ranges from structural chemistry to applied chemical problems such as those which exist in analytical chemistry. The demands that technological society has placed upon analytical chemistry have been met to an appreciable extent by advantage being taken of the unique properties of metal chelates. In general, however, chelating agents suffer the disadvantage of being unselective* in their reactions with metal ions, and one of the main objectives of analytical chemistry has been the development of new chelating agents which exhibit improved selectivity. To this end, the measurement of stability constants

---

*As defined by Feigl (4), a selective reagent gives an analytical reaction with a limited number of metal ions and a specific reagent gives an analytical reaction with only one metal ion.

has been important in understanding the factors that affect the solution stability of metal chelates and to some extent, therefore, the selectivity of chelating agents. This thesis is concerned with new 8-hydroxyquinoline chelating agents and the stability* of their metal chelates.

---

*The term "stability" refers to the solution stability of the metal chelate, as defined by the concentration equilibrium constant. The terms "stability constant" and "formation constant" are used interchangeably throughout.

HISTORICAL INTRODUCTION

Selectivity of Organic Analytical Reagents

The importance of organic reagents in chemical analysis is widely recognized. Chelating agents have been used extensively in qualitative(5), gravimetric(6), titrimetric(7), spectrophotometric(8), ion-exchange(9) and solvent-extraction(10) methods, and, to a lesser degree, in other methods too numerous to list.

Despite the usefulness and versatility of organic reagents, they suffer certain disadvantages, the most important of which is the general lack of selectivity in their reactions with metal ions. A considerable amount of research has been directed towards the determination of the factors which govern selectivity, in the hope that an understanding of these factors would assist in the design of new specific ligands. In essence, these studies have been focused on factors that affect the solution stability of metal complexes since, in general, the selectivity of a ligand is related to differences in the stabilities of its metal complexes. The factors which govern metal-complex stability have been discussed extensively(1-3, 11-17) and only the more important ones are listed here. These are:

(1) the basicity of the ligand donor atoms;

(2) the nature of the ligand donor atoms, e.g., whether hard base (oxygen) or soft base (sulfur);

(3) the number of donor atoms per ligand (i.e., the number of chelate rings formed);

(4) the size of the chelate rings;

(5) steric effects resulting from suitably placed substituents in the ligand;

and (6) properties of the metal ion such as ionization potential, charge/radius ratio, hardness or softness as acids, ability to form π bonds, and preferred stereochemistry.

Although our knowledge of the factors governing the stability of metal complexes has been greatly increased, its application to the discovery of selective reagents has not been very rewarding. Perhaps the most fundamental application has been in the design of ligands which show selectivity between groups of metal ions. It is well known, for example, that ligands with sulfur donors are selective for metal ions with definite class B (soft) behaviour (e.g., Pt(II), Au(I), Hg(II)), while ligands with oxygen donors are selective for class A (hard) metal ions (e.g., Zr(IV), Th(IV), Hf(IV)). Oxalic acid and dithizone are examples of ligands selective for class A and class B metal ions, respectively. The number of highly selective reagents, however, is disappointingly small. Calcichrome(18), 2,9-dimethyl-1,10-phenanthroline(19) and dimethylglyoxime(20) are among the few highly selective reagents.

A class of organic compounds which has received particular attention with respect to stability and selectivity studies is the 2-substituted 8-hydroxyquinolines*, of which the ligands described in this thesis are members.

---

*Hereafter, this ligand will be referred to as oxine.

2-Substituted Oxines

Oxine is a familiar reagent in analytical chemistry, with applications in gravimetric, titrimetric, spectrophotometric, fluorimetric, polarographic and amperometric analysis, and in solvent extraction and chromatography(21). Consequently, the oxine family of ligands has attracted considerable attention with respect to fundamental studies on stability and selectivity.

In 1944, Merritt and Walker(22) reported that 2-methyl-8-hydroxyquinoline* resembled oxine in giving water-insoluble chelates with a large number of metal ions. A notable exception was its lack of reaction with Al(III). This unexpected selectivity was exploited for the determination of Mg(II) and Zn(II) in the presence of Al(III). Merritt and Walker explained this lack of reactivity in terms of ligand-ligand steric interactions, and suggested that because of the increased size of 2-MeOx in relation to oxine the grouping of three large molecules around the small Al(III) ion would not be possible. Although this type of steric interaction should not prevent the formation of a 1:1 complex, Phillips and Price(23) observed that a 1:1 complex does not form to an appreciable extent in aqueous solution.

Irving and Pettit(24) suggested that the failure of 2-MeOx to form even a 1:1 species with Al(III) in aqueous solution is due to the formation of hydrolyzed species which are more stable than the sterically destabilized Al(III) chelate. Horton(25) apparently prepared the tris compound in molten 2-MeOx and Fernando _et al_.(26-28) prepared a number of Al(III) chelates of 2-MeOx in non-aqueous medium. These later workers suggested that the absence of a reaction in aqueous solution was caused by steric interactions between

---

* Hereafter, this ligand will be referred to as 2-MeOx.

the methyl substituent and the adjacent coordinated water molecules of the metal ion. Recently, Cardwell(29) showed that the tris complex can be precipitated from basic aqueous solutions in the presence of acetate ion. This secondary ligand prevents hydrolysis of the Al(III) as the pH is raised to promote metal-chelate formation.

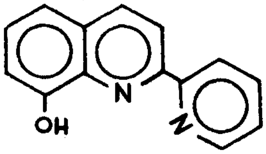
In studies with divalent ions of the first-row transition metal ions, Johnston and Freiser(30) found that the chelates of 2-MeOx were less stable than those of oxine (despite the higher pK_a(NH) value of 2-MeOx), and attributed this observation to steric interactions involving the 2-substituent. This effect was later found to be general with alkyl substituents(31). With the larger substituents such as 2-n-butyl and 2-phenyl, and also with the 2-thienyl group(32), a reversal in the magnitude of the stability constants was found, i.e., $\log K(ML_2) > \log K(ML)$. This behavior is unusual since stepwise stability constants normally decrease in a regular manner with increasing coordination, owing primarily to statistical, electrostatic and steric effects(11, 33). The effect was interpreted in terms of steric interaction between the 2-substituent of the first incoming ligand and an adjacent coordinated water molecule, resulting in a distortion of the aquated metal chelate and a consequent reduction in log K(ML).

In two further studies, oxine ligands with 2-substituents containing a potential donor atom were synthesized to determine whether the substituent acts as a sterically-hindering or complexing group. In the first study(34), the ligands examined were 4-amino-5-hydroxyacridine, 4,5-dihydroxyacridine and 4,5-diaminoacridine. These rigid ligands were treated as 2-substituted oxines. In the second study(35), these ligands were 2-aminomethyl, 2-carboxaldehyde and 2-hydroxymethyl-oxine in which the 2-substituents containing the potential donor atom are non-rigid, relative to the acridine ligands. In general, the 2-substituent steric effect was observed in both studies. In addition, chelate-ring strain proved to be an important factor in the formation of complexes with the acridine ligands. Of all the ligands examined, only 2-aminomethyl-oxine was found to yield terdentate complexes of high stability, although some question concerning the interpretation of this work exists and is dealt with under RESULTS AND DISCUSSION.

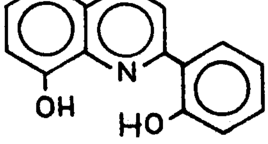
Purpose of this Research

Since few of the 2-substituted oxines described in the preceding section form terdentate chelates, synthesis of further oxines with potential donor 2-substituents was felt warranted. From the previous studies it is evident that terdentate behavior is least favored by rigid ligands such as substituted acridines. Accordingly, the following ligands were prepared and studied: 2-(2'pyridyl)-8-hydroxyquinoline (I), 2-(2'-hydroxyphenyl)-8-hydroxyquinoline (II) and 2-hydrazino-8-hydroxyquinoline (III)*.

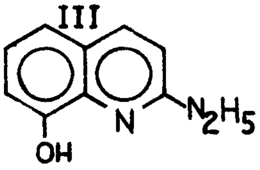
I



II



III



---

*Hereafter, these ligands will be referred to as 2-PyOx, 2-PhOx, and 2-HyOx, respectively.

The study concerned:

(1) the nature of the complexes formed with Al(III), Co(II), Ni(II), Cu(II), Zn(II) and Cd(II) in solution;

(2) the determination of the solution stability of these metal chelates;

and (3) comparison of the results of (1) and(2) with data for related ligands.

Potentiometric Determination of Formation Constants

The solution stability of a metal complex is concerned with the position of equilibrium in a reaction of the form

$$M + nL \rightleftarrows ML_n \qquad [1]$$

in which charges are omitted for clarity. The overall formation constant for this reaction is expressed as

$$\beta(ML_n)^* = [ML_n]/([M][L]^n) \qquad [2]$$

Since reaction [1] rarely occurs in one step, stepwise formation constants such as

$$K(ML_n) = [ML_n]/([ML_{n-1}][L]) \qquad [3]$$

can also be written.

These constants are concentration equilibrium constants and the square brackets signify concentration units of moles/liter. More strictly, the solution stability is defined by the thermodynamic equilibrium constant in which unitless terms replace the concentration terms. Because

---

* The species in the brackets indicate the complex to which the constant refers.

of the difficulty in obtaining activity coefficients for many of the species involved in metal-complex equilibria, thermodynamic constants have been obtained in only a limited number of studies. Since most stability constants are obtained for purposes of comparison, concentration constants are adequate, providing the experimental conditions are controlled properly. The parameters requiring regulation are temperature, ionic strength, and solvent composition.

The most accurate and widely applicable technique for the determination of formation constants is the pH-potentiometric method. This method is limited to protonic ligands and is based on the detection and measurement of the protons which are liberated when chelation occurs as shown below:

$$M \ +nHL \rightleftarrows ML_n \ +nH^+$$

The concentration of free or uncomplexed ligand is pH dependent and thus a large range of experimental ligand concentrations can be obtained. Most of the data collected by this method has been interpreted on the basis of the model originally suggested by Bjerrum(36), i.e., that the formation of complexes occurs in step-wise fashion with the stabilities of the various species characterized by a series of equilibrium constants. The methods most frequently used for evaluating stability constants involve relationships between the formation function, $\bar{n}$, and the free ligand concentration. The calculations are tedious and, in general, approximation techniques are employed.

The increasing availability of high-speed digital computers has now made it feasible to use standard objective treatments such as least squares adjustment(37). This technique can be applied to simple systems such as the graphical procedure given by Irving and Rossotti(38) for the determination of K(ML) and K(ML_2). Recently, the tendency has been to develop more general programs of greater mathematical complexity(39). The program SCOGS (Stability Constants of Generalized Species)(40), which has been used extensively in this work, is an example. It allows both the calculation of formation constants and the determination of the types of metal-complex species present. In spite of its advantages, this program has one serious limitation. Where reliable constants are to be calculated for many coexistent species, it is essential to have experimental data of the highest attainable accuracy, covering a wide range of experimental concentrations of ligand, metal-ion and acidity. If these conditions are not satisfied, the indiscriminant use of such methods may yield erroneous species or inaccurate constants.

The experimental procedure used in this study for the potentiometric titration was that of Calvin and Wilson(41). These workers used an acidified aqueous dioxane solvent system containing the ligand and metal ion for titration with a solution of standard base. The mixed solvent was used to overcome difficulties associated with the limited aqueous solubility of many metal chelates. This solvent does introduce problems, however, with respect to the measurement of pH and to the formation of ion pairs.

The hydrogen-ion concentration can be measured with a glass and calomel electrode assembly. Since activity coefficients depend on the solvent composition, no significance can be attached to absolute pH values obtained in aqueous organic solvents when the pH-meter has been previously calibrated with an aqueous buffer. Also, when a reference electrode, which is designed for use in pure aqueous solvents, is employed in a partial non-aqueous medium, the value of the liquid-junction potential is different from that in a purely aqueous solution. For these reasons, this electrode assembly must be calibrated.

Van Uitert and Haas(42) have shown that the glass electrode functions linearly, with respect to the hydrogen electrode, over the pH range of 1.5 to 11 in aqueous dioxane solutions containing up to 75% v/v dioxane. Consequently, the glass electrode can be used as a concentration probe. The correction which is necessary to convert the pH-meter reading into $-\log[H^+]$ is given by

$$\log U_H = -\log[H^+] - pH$$

where U_H = correction factor,

$[H^+]$= equilibrium concentration of hydrogen ion,

and pH = meter reading.

U_H is a function of the solvent concentration and ionic strength, and is constant if these factors are controlled.

Although the potentiometric method is widely applicable, it has a number of limitations. The upper pH limit is determined by the hydrolysis of the metal ion. For most divalent transition ions, the pH range of hydrolysis in 50% v/v aqueous dioxane is 6-7. The lower pH limit depends on the ability of the glass electrode to detect the protons liberated on chelation in a solution which is already about 0.005M in perchloric acid. This pH value is approximately 2.5. If the stability of a complex is either very large or very small, chelation will occur outside this pH range, and the potentiometric method fails. Finally, meaningful measurements cannot be made if precipitation of the metal chelate or other metal-containing species occurs during the titration.

EXPERIMENTAL

Apparatus

Calibrated volumetric ware was used throughout this work whenever this was appropriate.

Infrared spectra were recorded with a Beckman IR-5 infrared spectrophotometer (Beckman Instruments, Inc., Fullerton, California). Spectra in the visible region were recorded with a Bausch and Lomb Spectronic 600 spectrophotometer (Bausch and Lomb, Rochester, N.Y.). Nuclear magnetic resonance spectra were recorded with either an A-60 or a HA-100 Varian nuclear magnetic resonance spectrometer (Varian Associates, Palo Alto, California). Polarograms were recorded on a Polarecord E 261 polarograph (Metrohm, Switzerland). A dropping mercury cathode and a saturated calomel reference electrode were used.

The potentiometric titration apparatus for the determination of formation constants consisted of a titration cell, a pH-meter, two 10-ml semi-microburettes and a constant temperature bath. The pH-meter was a Radiometer Model PHM4C instrument (Radiometer, Copenhagen, Denmark) equipped with a Beckman saturated calomel and a Beckman E3 glass electrode (low sodium-ion error).

Reagents

The common chemicals used were either of analyzed grade or of a purity suitable for the purpose intended.

Carbonate-free sodium hydroxide ($\sim$ 0.1M) was prepared by a standard procedure(43) and stored in a 4-litre polyethylene bottle equipped with a tube of ascarite for removal of carbon dioxide. The background salt concentration (sodium perchlorate) in this solution was 0.2 molar and the molarity was determined by titration of standard potassium acid phthalate to a phenolpthalein endpoint.

The perchloric acid (∿ 0.01 M, 0.19 M in sodium perchlorate) used in the potentiometric titrations was standardized by potentiometric titration of 50-ml aliquot portions with standard sodium hydroxide solution.

1,4-Dioxane (Fisher Certified Grade) was purified by refluxing over sodium metal for at least 36 hours prior to distillation through a one-metre column packed with glass helices. The fraction boiling in the range 100.5 to 101.0°C was collected as required and used within 24 hours.

The metal-ion solutions (∿ 0.01 M, 0.18 M in sodium perchlorate) used for potentiometric titrations were prepared from the perchlorate salts and standardized with EDTA by accepted procedures(44). With appropriate dilution factors, these solutions also served for polarographic and spectrophotometric studies of the metal complexes.

Synthesis of Ligands

Reaction sequences for the synthesis of 2-PyOx, 2-PhOx and 2-HyOx are given in Figures I, II and III, respectively. The detailed procedures for the syntheses are given in APPENDIX I, together with data from elemental analysis. The ligands were further characterized by infrared, proton nuclear magnetic resonance and mass spectrometric analysis, and

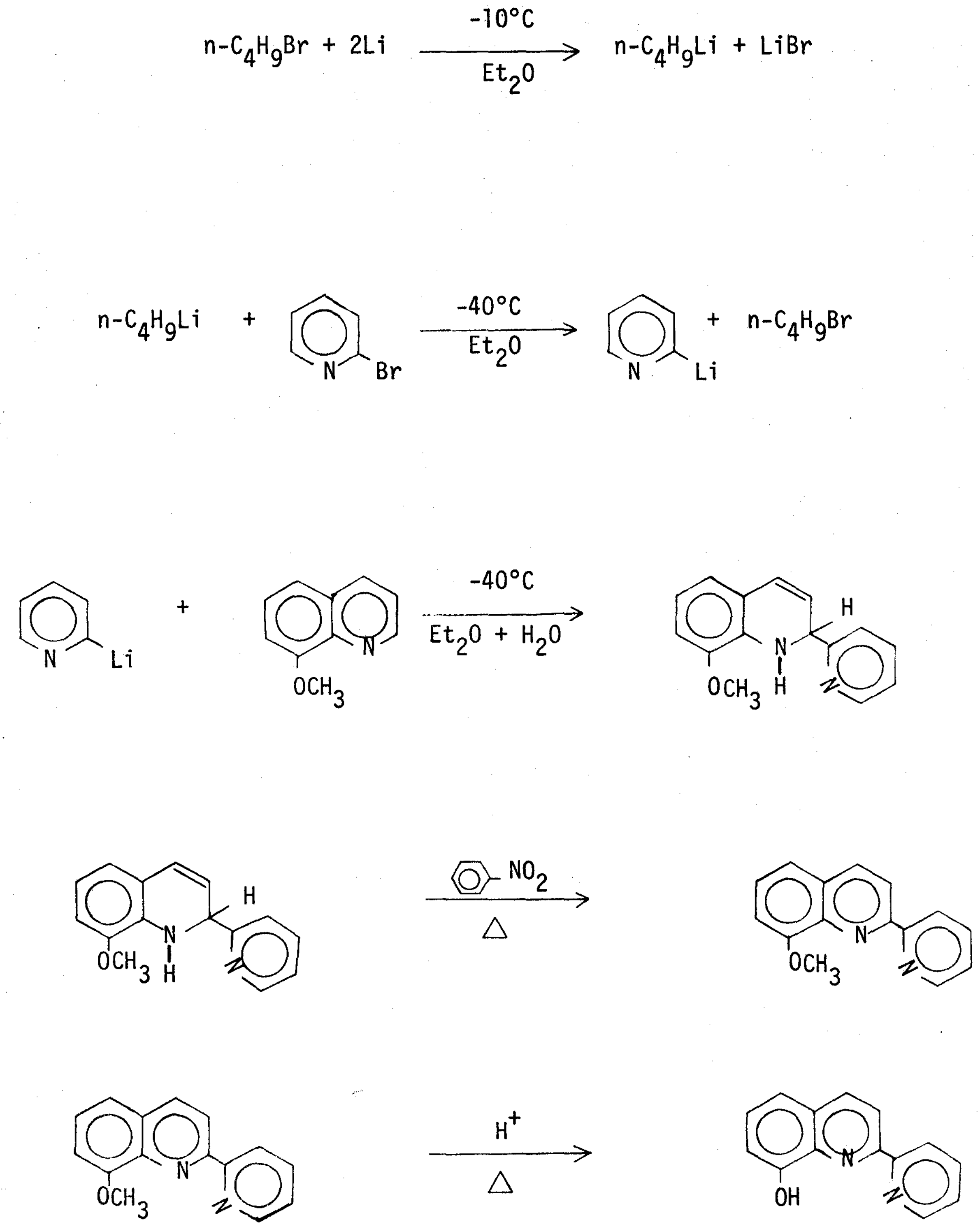


Figure 1. SYNTHESIS OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE.

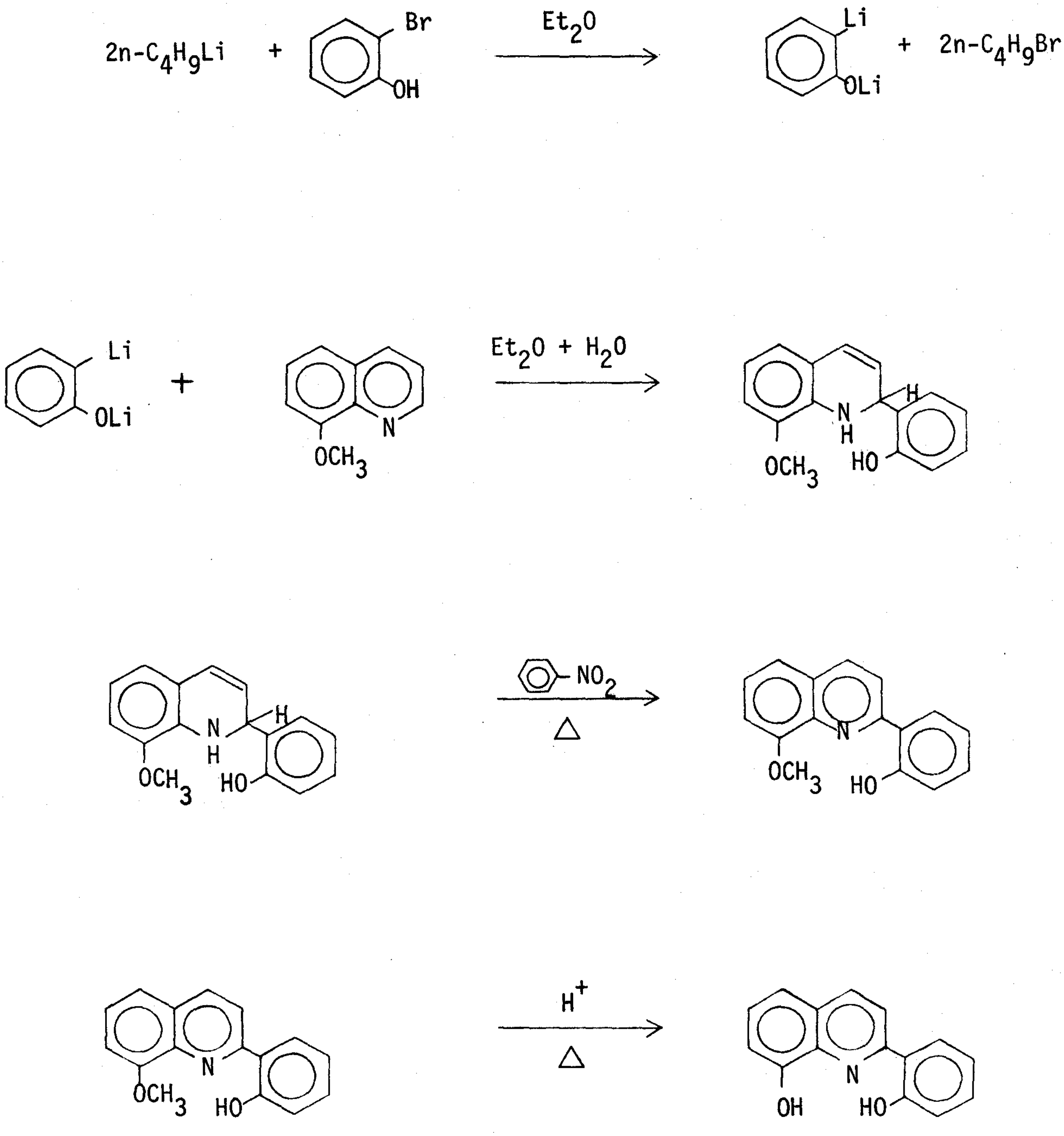


Figure 2. SYNTHESIS OF 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE.

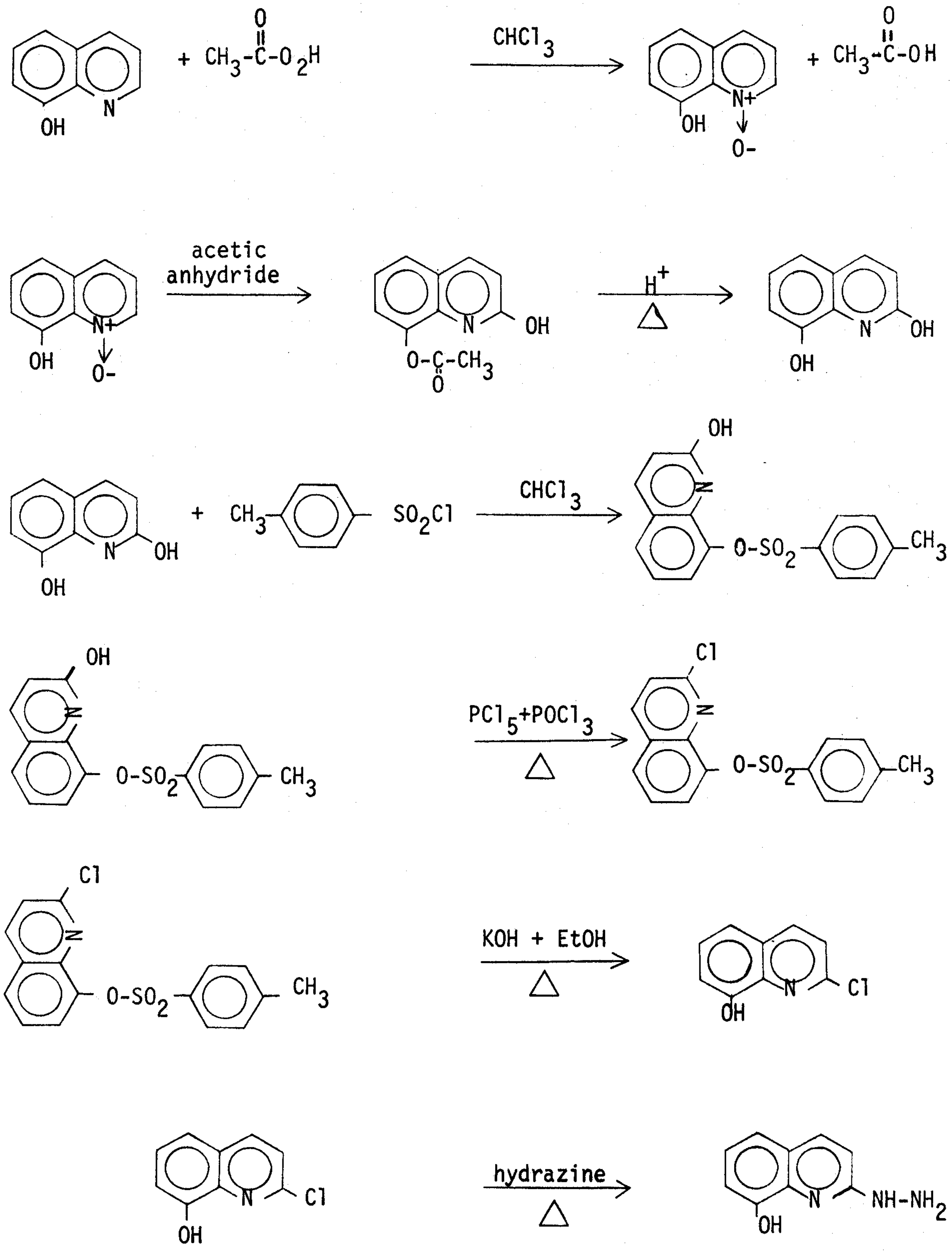


Figure 3. SYNTHESIS OF 2-HYDRAZINO-8-HYDROXYQUINOLINE.

the data are presented under RESULTS AND DISCUSSION.

Preparation of Solid Chelates

2-PyOx Chelates with Composition ML_2. Eight milliliters of a 0.01 M metal-ion solution (Ni(II), Co(II) and Zn(II)) were added slowly to a stirred solution consisting of 50 mg of 2-PyOx, 60 ml of water, 15 ml of dioxane and 5 ml of 0.1 M perchloric acid. The temperature was maintained at 70°C and 0.1 N sodium hydroxide was added slowly until a pH value of 7-8 was obtained. After a 45-minute digestion period, the precipitate was filtered, washed with water and dried under vacuum for 2 hours.

2-PyOx/Ni(II) Chelate with Composition $NiLHLClO_4$. This chelate was prepared by two methods. In the first method, the procedure was the same as that above except that the amount of ligand and metal ion was doubled and the pH was adjusted to 2.5. In the second procedure, 90 mg of the 2-PyOx/Ni(II) chelate, NiL_2, were placed in 15 ml of dioxane and the resulting mixture was stirred. An equivalent molar quantity of 1 M perchloric acid was added by means of a microburette. Eight milliliters of alcohol were added to increase the solubility of NiL_2 and the resulting mixture was allowed to stir for three days. The precipitate was then filtered off and dried under vacuum for several hours at room temperature.

The microanalytical data for these solid 2-PyOx chelates are given in TABLE VII of RESULTS AND DISCUSSION.

2-PhOx and 2-HyOx Chelates. The preparation of solid chelates of 2-PhOx and 2-HyOx was unsuccessful. A procedure similar to that for the preparation of the ML_2 chelates of 2-PyOx was used, but it was found that

2-HyOx decomposed under these conditions and that the 2-PhOx chelates were impure, most likely due to hydrolysis and/or contamination by excess ligand.

Reactivity of Ligands with Metal Ions

Spot tests with 24 metal ions were performed to determine the general reactivity of the ligands towards metal ions over a wide range of pH. The procedure was as follows: one drop of a metal-ion solution (∼0.02 M) was added to a test tube containing one ml each of water and a buffer solution.* Fifteen drops of an alcoholic reagent solution (0.5% w/v) were added slowly while the test tube was agitated vigorously. Three milliliters of chloroform were then added and the test tube was shaken again. The appearance of a color in either the aqueous or non-aqueous phase or the formation of a precipitate was noted. Both metal-ion and reagent blank tests were made under similar conditions. The results are given in TABLES II, III and IV in RESULTS AND DISCUSSION.

Ultraviolet-Visible Absorption Spectra

Procedure. The following procedure was used to obtain spectra of 2-PyOx, 2-PhOx, and 2-HyOx and chelates as a function of pH, for mole-ratio studies of 2-PyOx and 2-PhOx chelates, and for the determination of the acid-dissociation constant of the phenolic-OH group of 2-PhOx. The required amount of a dioxane solution of reagent (10^{-4}M) was pipetted

---

*Hydrochloric acid was used for pH -0.3 and 1.0,phosphate buffers (0.02 M in total phosphate) for pH 2.6 and 7.0, an acetate buffer (0.2 M in total acetate) for pH 5.0, and an ammonia buffer (0.2 M in total ammonia) for pH 10.

into a 50 ml volumetric flask and then diluted with dioxane to 25.0 ml. The desired pH was obtained with hydrochloric acid (for solutions having a pH of 3 or less), or sodium hydroxide (for solutions having a pH of 12 or greater) or with 5 ml of a phosphate solution (0.05 M in total phosphate) for pH values of 4, 7, 8 and 11, or 5 ml of an acetate solution (0.05 M in total acetate) for pH values of 5 and 6. When possible, the ionic strength was maintained at 0.1 by addition of 1.0 M sodium perchlorate. If required, the metal-ion solution (∿ 0.1 M) was then added and the total volume of water was adjusted to 25.0 ml. After equilibration at 25°C, the solution was diluted to the mark with 50% (v/v) aqueous dioxane. Blank solutions were prepared for each sample and the spectra were recorded using 1-cm fused-silica absorption cells. The pH of each solution was measured as soon as the spectrum was recorded.

The spectra of the ligands are given in APPENDIX IV. The spectra of the 2-PyOx chelates and data from the mole-ratio and acid-dissociation studies are presented in RESULTS AND DISCUSSION.

Acid-Dissociation Constant of 2-PhOx. Since the pK_a value for the phenolic-OH group of 2-PhOx is too high for potentiometric determination, a spectrophotometric technique was employed. Absorption measurements were taken at two different wavelengths (291 and 302nm) and sets of two simultaneous equations were solved to give the acid-dissociation constant. Details of the calculations are given in APPENDIX II.

Potentiometric Determination of Acid-Dissociation Constants

Titration Procedure. The potentiometric titrations were performed in a 250-ml jacketed cell. The contents of the cell were maintained at a constant temperature (± 0.1°C) by circulation of water at the desired temperature (25 or 60°C) between the walls of the cell. The cell was covered with a lucite plate equipped with holes for two semi-microburettes, reference and glass electrodes, nitrogen gas inlet and outlet tubes and a thermometer. Nitrogen gas was bubbled through the solution during the titration except when titrations were done at 60°C.

Prior to each titration, the pH-meter was standardized at pH 4.01 and 7.00 with Harleco standard buffer solutions. A weighed quantity of the ligand (∼ 0.7 Millimoles) was placed in the titration cell along with 50.0 ml of standard 0.01 M perchloric acid and 50.0 ml of dioxane. The electrodes were allowed to equilibrate with this solvent system for 15 minutes before the titration was started. The titrant was a standard 0.1 M sodium hydroxide solution. After each addition of base, an equal volume of dioxane was added to maintain the 50% (v/v) composition of the solvent medium. The pH was then recorded. Electrode response was very rapid and drifting normally occurred only with precipitation or in solutions of low buffer capacity.

Back-titration with standard 0.1 M perchloric acid was attempted to obtain a second set of data, but it was found that the glass electrode exhibited hysteresis effects. Karlberg and Johansson (45) have recently reported similar hysteresis behavior of the glass electrode in nonaqueous media. Representative titration data are given in APPENDIX III.

Volume Corrections. When dioxane and water are mixed, the final volume is not equal to the sum of the initial volumes. At 25.0°C, there is a net volume shrinkage and the exact percent volume change for initially equal volumes of dioxane and water was determined as follows: twenty-five milliliters each of water (0.2 M in sodium perchlorate) and dioxane were placed in a clean, dry calibrated 50-ml volumetric flask, mixed well without inverting the flask and then equilibrated at 25.0°C. The volume of 50% (v/v) aqueous dioxane required to bring the solution in the volumetric flask to the mark was measured with a 1-ml graduated pipet. Three determinations gave a mean volume contraction of 1.8 ± 0.1%. This value is in agreement with the previously determined value of 1.6%(46).

When the temperature of a 50% (v/v) aqueous dioxane solution is raised from 25°C to 60°C a net expansion occurs. To measure this expansion, 50-ml volumetric flasks were filled to the mark with water at 25.0°C and then equilibrated at 60°C. The expansion coefficient of water was used to calculate the volume change of the flasks and this was used for the computation of the percent expansion when 50% aqueous dioxane was used. When the initial shrinkage at 25°C was taken into consideration, the net overall expansion was found to be 1.0 ± 0.1%.

Electrode Calibration. As previously stated, before the glass electrode can be used as a hydrogen-ion concentration probe in aqueous dioxane, it must be calibrated to allow conversion of the pH-meter reading into hydrogen-ion concentration. The correction factor, log U_H, was determined at 25 and 60°C by titration of 50.0 ml of 0.01 M perchloric acid in 50.0 ml of dioxane with standard base. Hydrogen-ion concentrations were calculated assuming complete dissociation of the acid. U_H was not determined for titration points beyond 80% neutralization due to the large errors in the pH-meter readings in this region. For a 50% aqueous dioxane solution with an ionic strength of 0.1, the values obtained for log U_H were -0.070 ± 0.008 at 25°C and +0.065 ± 0.005 at 60°C. Irving and Mahnot have determined a number of values for log U_H in 50% aqueous dioxane. They record a value of -0.082 ± .006 at 25°C(47).

Calculation of pK_w. Values of pK_w were calculated with data from the same titrations used to determine U_H. For each titration point past the equivalence point, the value of $[OH^-]$ was determined from the normality of the base, and the pH-readings were corrected for both the sodium-ion error and the solvent system (i.e., log U_H) to give $[H^+]$. The values of pK_w were subsequently found to be 15.33 ± 0.02 at 25°C and 14.28 ± 0.02 at 60°C. The value of pK_w at 25°C agrees with that found by Billo and Corsini(32).

Incomplete dissociation of sodium hydroxide in the solvent medium will give rise to errors in the calculation of pK_w. Conductivity studies(48,49) have shown that partial association does occur in 45% dioxane. The extent of this association is small, however, and since the reliability of the dissociation constant for sodium hydroxide is poor, there is likely very little to be gained by correcting for ion-pair formation.

Calculation of Acid-Dissociation Constants. A detailed derivation of the equations used appears in APPENDIX II. A computer program, designed to perform three main operations, was developed for the calculation of acid-dissociation constants. The first operation was the calculation of the variable $\bar{p}$, the average number of protons bound to each ligand molecule (APPENDIX II). The values of $\bar{p}$ were then plotted against pH. Second, a plot of pH versus volume of base was made. Third, the value of K_a was calculated for each point using the experimental concentrations of hydrogen ion, dissociated acid and undissociated acid. Since unreliable pH-meter readings are obtained in regions of low buffer capacity, the values of K_a were averaged only over the points where $\bar{p}$ varied from 0.2 - 0.8, 1.2 - 1.8 or 2.2 - 2.8, depending upon the dissociation under consideration. The graphs of $\bar{p}$ versus pH and pH versus ml base were used to check the extent to which the stepwise dissociation equilibria overlapped. This check was necessary because a different set of equations would be required if the equilibria did overlap. The graphs were also useful for the rejection of spurious data.

The acid-dissociation constants for the ligands are given in TABLE I, RESULTS AND DISCUSSION, and representative data used in the calculations are given in APPENDIX III.

Potentiometric Determination of Chelate Formation Constants

Procedure. The procedure was the same as that used in the determination of acid-dissociation constants, except for the addition of 5.00 ml of a standard 0.01 M metal-ion solution and 5 ml of dioxane, to maintain the solvent composition. The hydrolysis curve for each metal ion was obtained by titration in the absence of ligand. A minimum of three titrations were performed for each metal-ion/ligand system, usually in the molar ratios of 5:1, 2:1 and 1:1[*]. Representative data are given in APPENDIX III.

Calculations. The full derivations of the equations used are in APPENDIX II. Four different computer programs were used for the calculation of formation constants. The appropriate program was determined by the ligand and metal ion under consideration (see below). The initial calculations involved three steps. In order, these were:

(1) calculation of $\bar{n}$, the average number of ligands bound per metal ion;

(2) construction of a plot of $\bar{n}$ versus pL, where pL is the negative logarithm of free ligand concentration, and of a plot of pH versus volume of base;

and (3) elimination of erroneous data and determination of the pH range for calculation of formation constants.

Oxine/Divalent Metal-Ion Systems. The formation constants for these systems were calculated in two steps. First, initial values of the constants were obtained from a linear least squares procedure. Second, these constants were refined by an iterative least squares procedure similar to that described by Unwin *et al.*(50). An estimation of the goodness of fit was obtained by calculating $\bar{n}$ values with these constants and by comparing the results with the experimental $\bar{n}$ values.

Oxine/Trivalent Metal-Ion System. Approximate values of the formation constants were obtained from the $\bar{n}$ curve and the final values were calculated by an iterative least squares procedure similar to that

---

[*] Unless otherwise indicated, ratios will always be written in the order moles of ligand:moles of metal ion.

used for the divalent metal-ion complexes. An estimation of the goodness of fit was obtained, as described above.

2-PhOx/Divalent Metal-Ion Systems. Since only 1:1 species were formed with this ligand, it was possible to calculate the formation constant for each point of the titration where chelation occurred (see APPENDIX II). The final value of K(ML) was taken as the mean value obtained in the $\bar{n}$ range of 0.2 - 0.8.

2-PyOx and 2-HyOx/Divalent Metal-Ion Systems. The program SCOGS$^{(40)}$ was corrected and revised (see APPENDIX II) to suit the 2-PyOx and 2-HyOx systems. By the program SCOGS, it is possible to calculate formation constants of species with the general formula $(M_1)_a\ (M_2)_b\ (L_1)_c\ (L_2)_d$ where a, b, c, and d have values of 0 to 10. The program uses initial estimates of the formation constants to calculate the concentration of all the species present. Then these formation constants are numerically differentiated and a second set of concentrations is obtained. The two sets of concentration data are used in a non-linear least squares procedure to obtain a new set of formation constants. This process is continued until the formation constants undergo no further significant changes.

The formation constants for these different systems are given in TABLES V, VI, VIII and IX in RESULTS AND DISCUSSION.

Polarography of Metal Complexes of 2-PyOx

Procedure. Initial experimental parameters were determined as follows: 5 ml of a 1 M sodium perchlorate solution, 5 ml of a buffer solution* and an aliquot of either a metal-ion solution, or of a dioxane solution of reagent, were added to a 50-ml volumetric flask. The total volumes of dioxane and water were each adjusted to 25.0 ml and the resulting solution was diluted to the mark with 50% aqueous dioxane. Approximately 25 ml of this solution were placed in a polarographic cell which was maintained at 25.0°C. A stream of nitrogen was passed through the solution for 15 minutes and 2 drops of a 0.2% (w/v) Titron X-100 solution were added. The polarogram was then recorded.

When polarograms were required for solutions being titrated potentiometrically, a 200-ml jacketed titration cell was used for both the titration and polarographic cell. The cell was covered with a lucite plate, through which holes were drilled for nitrogen gas inlet and outlet tubes, two reference electrodes, a dropping mercury electrode, a thermometer and two semi-microburettes. Eight drops of a 0.2% (w/v) Titron X-100 solution were added and nitrogen was passed through the solution for 20 to 30 minutes prior to recording the polarogram.

The results of this work are given under RESULTS AND DISCUSSION.

---

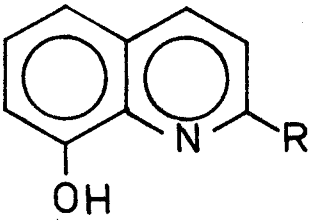
*These studies were done at pH values of 1, 2, 3, 4, 5, 6 and 8. The buffers used were the same as those used for the spectrophotometric studies.

RESULTS AND DISCUSSION

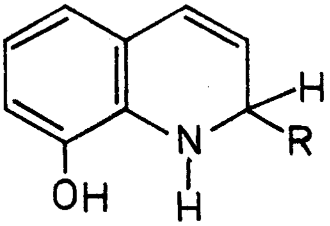
Synthesis and Characterization of Ligands

Organolithium compounds are commonly used in the preparation of 2-substituted oxines(I). The possible formation of a dihydro species (II) is referred to in the literature, but no clear evidence for its existence is given(21,31,51-57).

I



II



The formation of appreciable amounts of the dihydro species is relevent from a synthetic standpoint since a reduction in yield would occur. This reduction in yield could be offset, at least in part, by the addition of an oxidant. In several instances, the literature makes no reference to the addition of an oxidant in the synthesis of a 2-substituted oxines.

In the present work, initial attempts to prepare the previously unreported ligand 2-PyOx were made in the absence of an oxidizing agent. The resulting yields, based on 8-methoxyquinoline, were very low (2%). In a subsequent attempt, the use of nitrobenzene* as an oxidant increased the yield to approximately 20%. This ten-fold increase in yield strongly

---

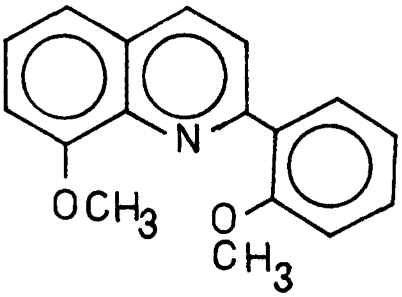
*Nitrobenzene is frequently used in the synthesis of several oxine derivatives to oxidize the resulting dihydro intermediate(21).

suggests the presence of the dihydro compound. The effect of other oxidizing agents was not examined.

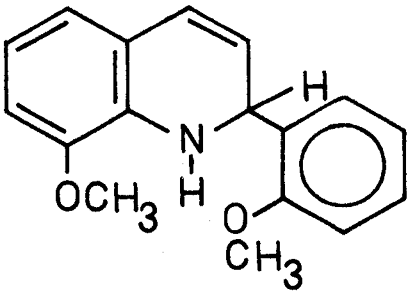
The synthesis of 2-PhOx has been reported previously(55). No oxidizing agent was used and the yield was not given. In the present investigation, a 1% yield was obtained when Towle's procedure(55) was followed. A yield of 16% was obtained when nitrobenzene was added to the reaction mixture.

Some evidence for the formation of the dihydro compound was obtained during the course of this work. Before Towle's procedure for the preparation of 2-PhOx was followed, o-lithiumanisole was used as a starting material. Reaction of this compound with 8-methoxyquinoline and subsequent hydrolysis yielded a viscous yellow oil. The mass spectrum of this oil indicated a molecular weight of 267 rather than 265 which would be expected for 2-(2'-methoxyphenyl)-8-methoxyquinoline(III). The value of 267 is consistent with the dihydro derivative (IV).

III



IV



The infrared spectrum of this compound (Spectrum 1, Figure 4) showed a strong N-H stretch absorption at ∿ 3400 cm^{-1}. This absorption was not observed after the original compound was refluxed with nitrobenzene

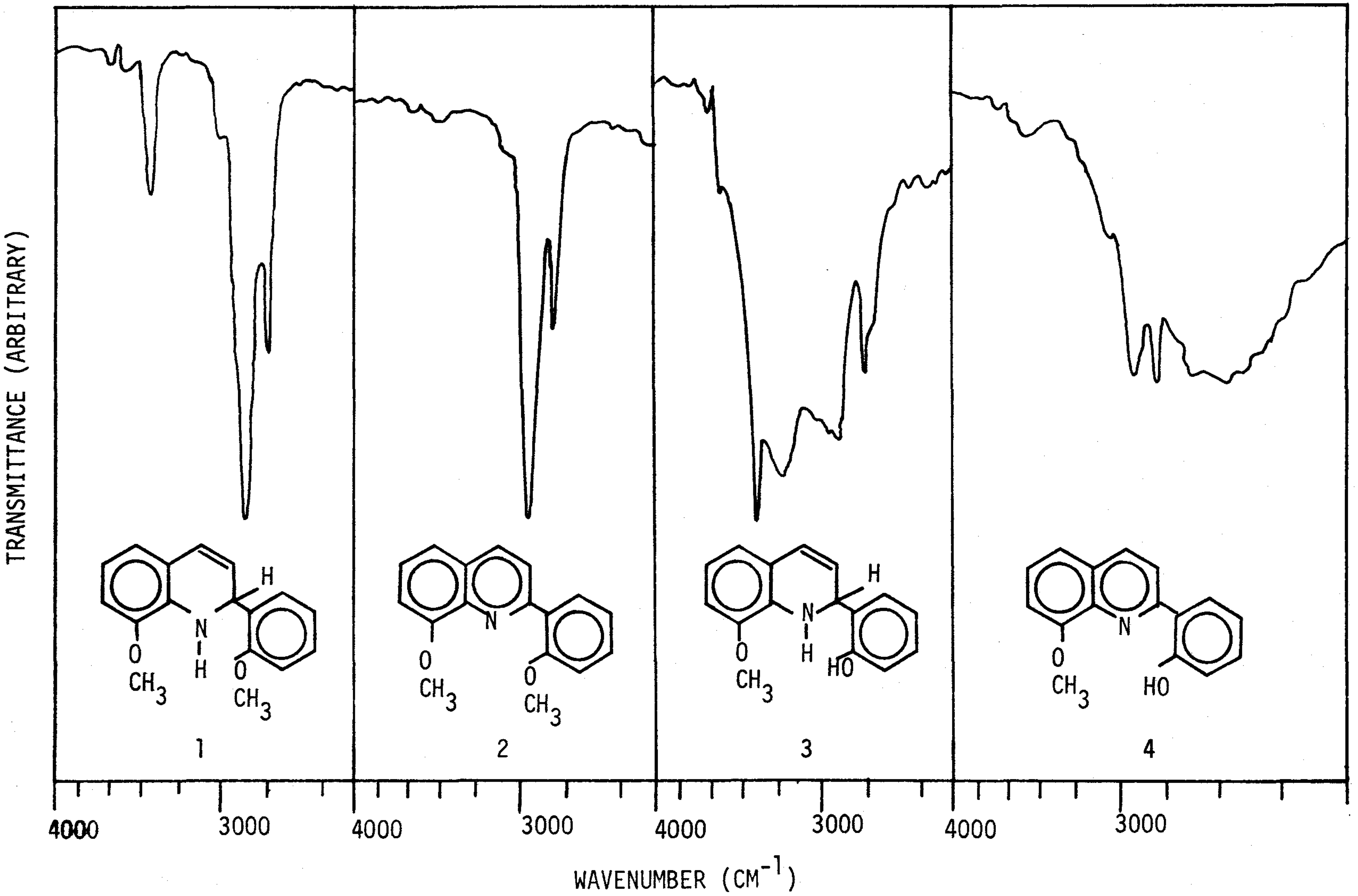


Figure 4. INFRARED SPECTRA OF INTERMEDIATES ISOLATED IN THE PREPARATION OF 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. All spectra recorded in $CHCl_3$, concentration = 25 mg/ml, 0.5-mm cells.

(Spectrum 2, Figure 4). When o-lithiumlithiumphenoxide was used in place of o-lithiumanisole, the resulting compound also displayed an absorption at ∿ 3400 cm^{-1} (Spectrum 3, Figure 4). In addition to this absorption, a possible O-H stretch absorption was observed at ∿3200 cm^{-1}. Oxidation with nitrobenzene removed the absorption at ∿ 3400 cm^{-1} and the O-H absorption was broadened and shifted to ∿ 2600 cm^{-1} (Spectrum 4, Figure 4). This broadening suggests an increase in hydrogen bonding, which could be attributed to the availability of the lone pair of electrons on the nitrogen atom.

The above discussion shows that dihydro intermediates can be important in the preparation of 2-substituted oxines when organolithium compounds are used as starting materials. The synthesis of 2-HyOx does not involve the use of an organolithium reagent and the procedure for this synthesis was not investigated.

Characterization of Ligands. 2-PhOx and 2-HyOx have been previously synthesized, but were not fully characterized (55,58). These two ligands and the new ligand, 2-PyOx, were characterized in the present study by infrared, proton nuclear magnetic resonance and mass spectroscopy and by elemental analysis.

The proton n.m.r. spectra of oxine, 2-PyOx and 2-HyOx are shown in Figures 5,6,7 and 8, respectively. It has been shown(59-61) that the spectrum of oxine can be interpreted as shown in Figure 5. This spectrum was used as a guide for the interpretation of the other spectra. The resonances at 15.6 and 10.6 ppm in the spectra of 2-PhOx have been assigned to H(9) and H(8), respectively, since 2-(2'-hydroxyphenyl)-8-

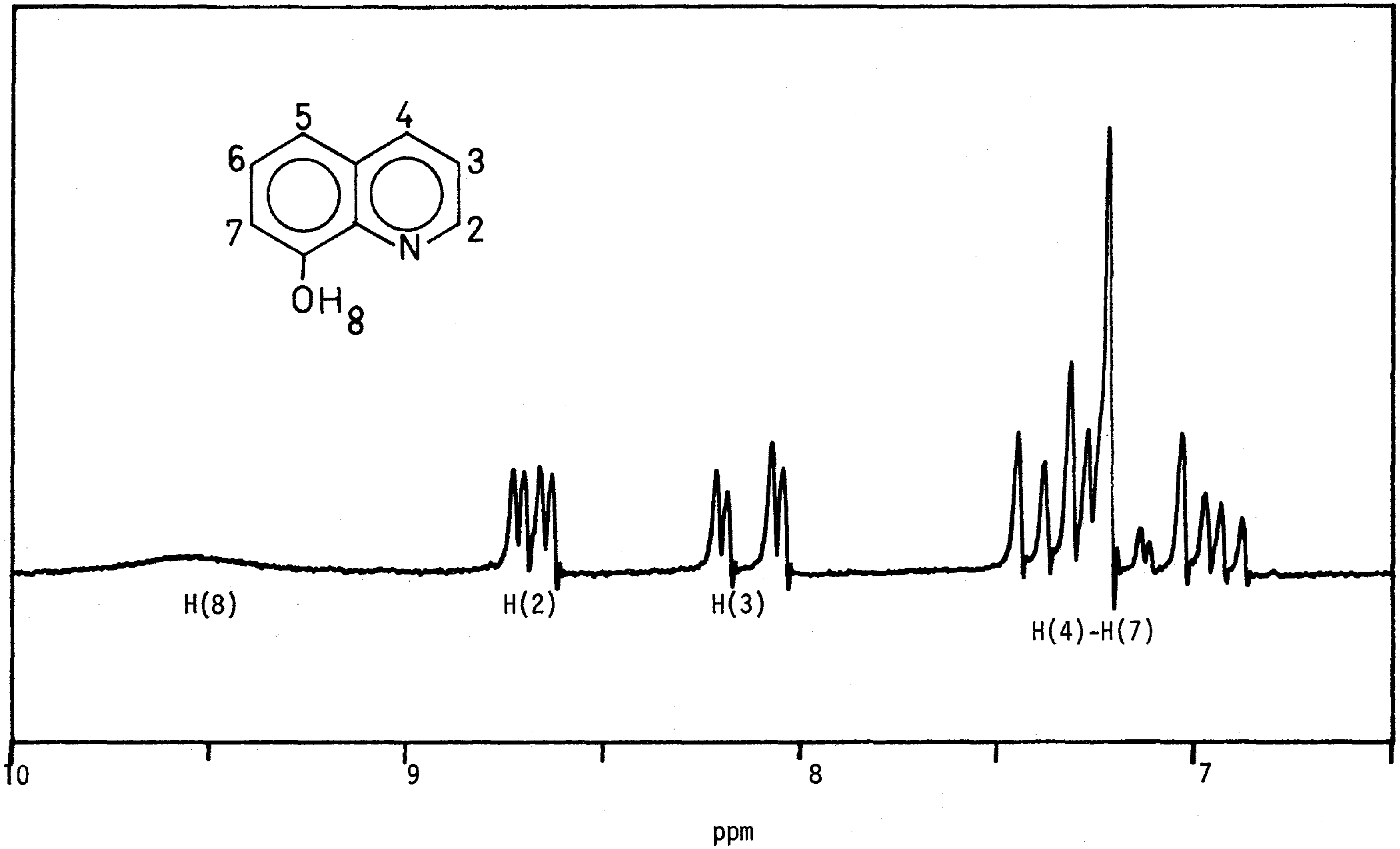


Figure 5. PROTON NUCLEAR MAGNETIC RESONANCE SPECTRUM OF 8-HYDROXYQUINOLINE. Solvent is dimethylsufoxide.

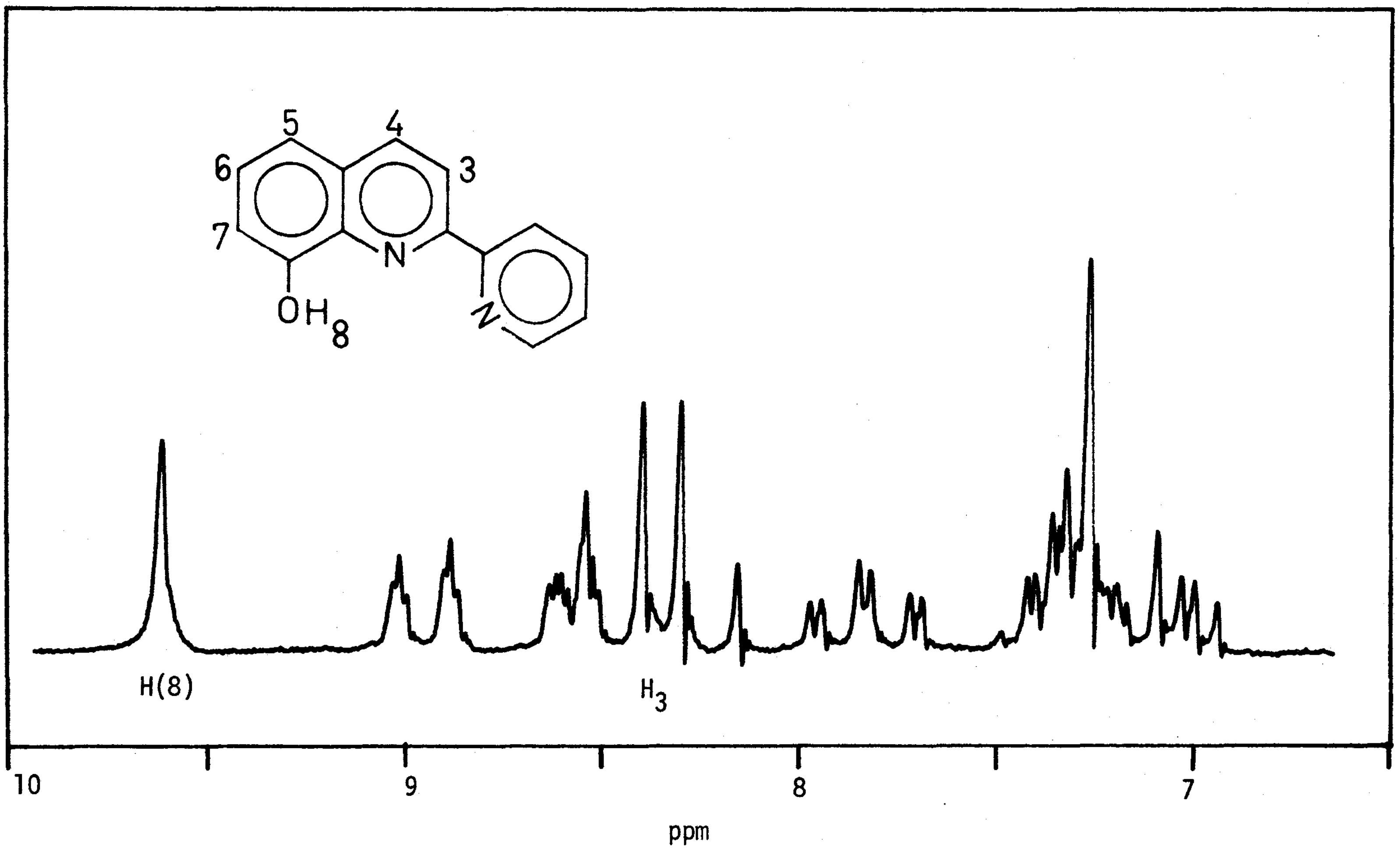


Figure 6. PROTON MAGNETIC RESONANCE SPECTRUM OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. Solvent is dimethylsulfoxide.

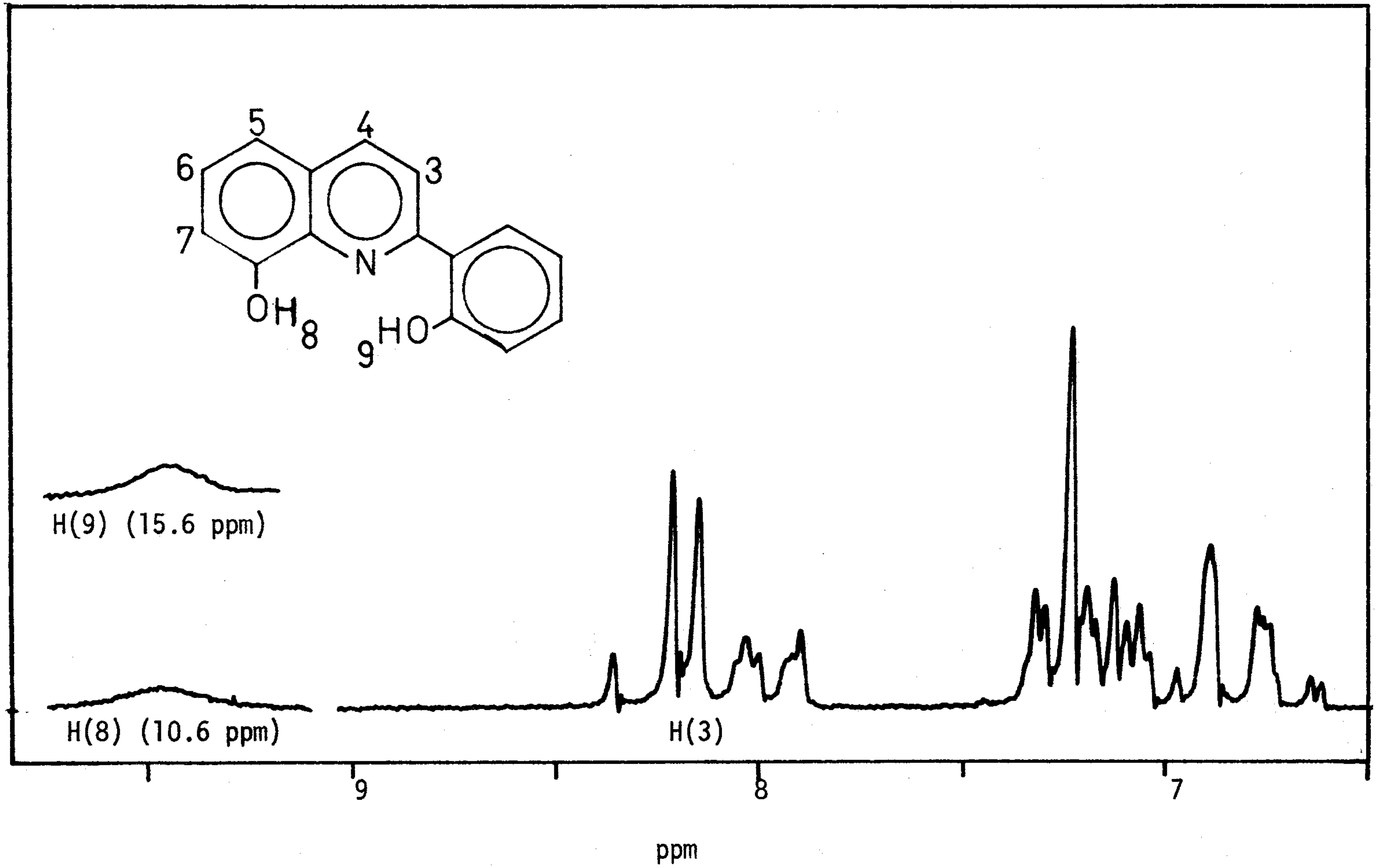


Figure 7. PROTON NUCLEAR MAGNETIC RESONANCE SPECTRUM OF 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. Solvent is dimethylsulfoxide.

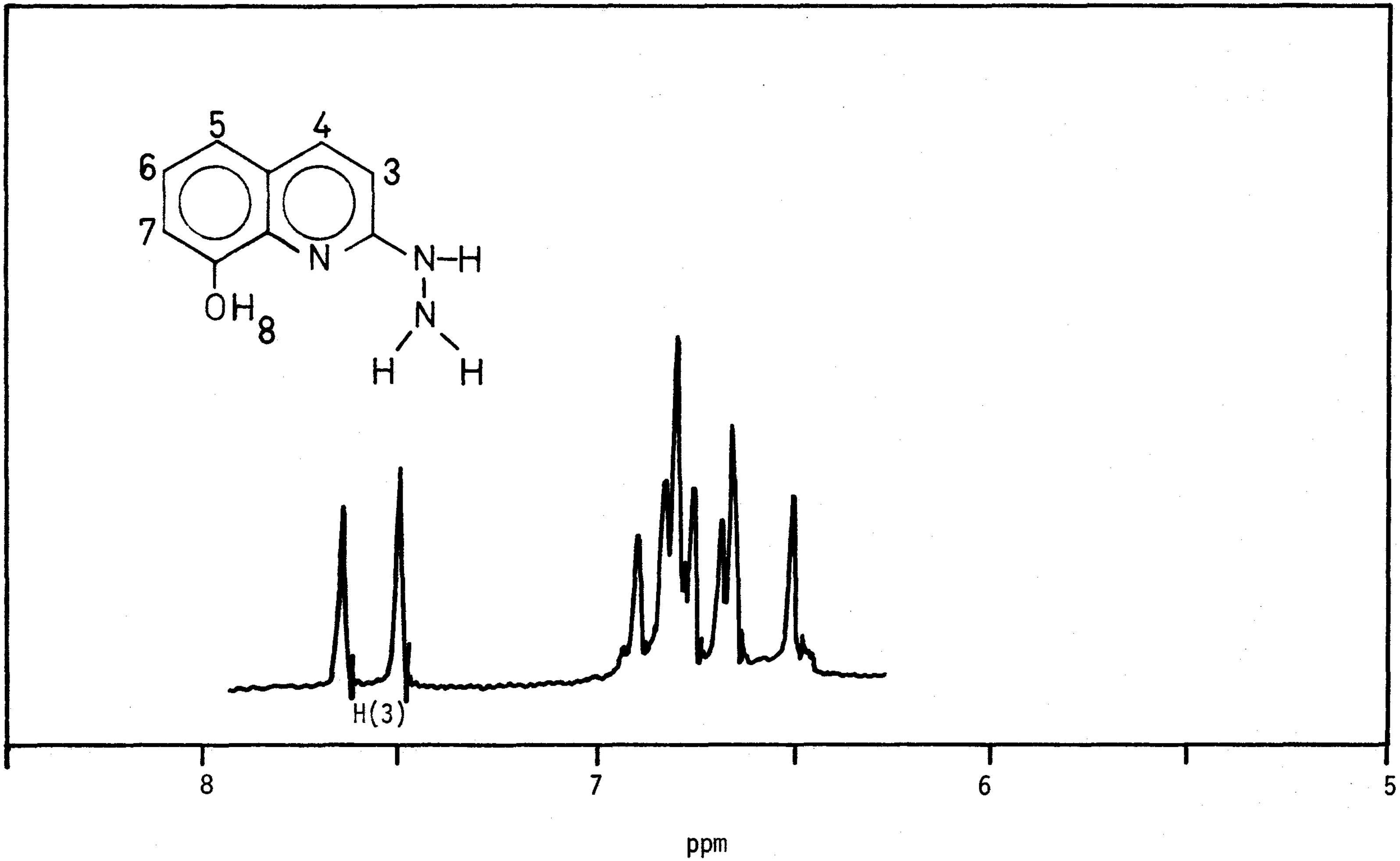


Figure 8. PROTON NUCLEAR MAGNETIC RESONANCE SPECTRUM OF 2-HYDRAZINO-8-HYDROXYQUINOLINE. Solvent is dimethylsulfoxide.

methoxyquinoline shows a single proton resonance at 16 ppm and not at 10.6 ppm (this spectrum has not been reproduced since it is similar to that of 2-PhOx). The integration for 2-PyOx and 2-PhOx accounts for all of the protons in these compounds, but for 2-HyOx, four proton resonances were not observed. The electric quadrupole moment of nitrogen could obscure the resonance of H(9), H(10) and H(11) of 2-HyOx (Figure 8) and possibly that of H(8) also, if this proton is hydrogen-bonded to the hydrazino group. A broad band near 2900^{-1}, indicative of a strongly hydrogen-bonded OH group, was observed in the i.r. spectrum (hexachlorobutadiene) of 2-HyOx.

Comparison of the proton n.m.r. spectrum of oxine with those for 2-PyOx, 2-PhOx and 2-HyOx shows that the H(2) resonance (quartet) in oxine is absent in these spectra. Thus, substitution has occurred in the 2-position of the quinoline ring for 2-PhOx, 2-PyOx and 2-HyOx.

The mass spectrum of 2-PyOx, 2-PhOx and 2-HyOx yielded parent-ion peaks which corresponded to the correct molecular weight for each compound. For the i.r. spectra of the ligands, and some of the intermediates in the synthetic procedures, attention was focused mainly on the regions 1200-1300 cm^{-1} (OCH_3) and 2900-3400 cm^{-1} (OH and NH). The spectra confirmed the presence of these functional groups in the ligands. The results of the elemental analysis (APPENDIX I) were consistent with the molecular formulas of the ligands.

The above results confirm the structures proposed for 2-PyOx, 2-PhOx and 2-HyOx in Figures 1-3 (EXPERIMENTAL).

Acid-Base Reactions of Ligands

The following section is concerned with the acid-base equilibria of 2-PyOx, 2-PhOx and 2-HyOx. An understanding of the acid-base properties of a ligand is important in studies relating to reactions with metal ions. This study proceeded in the following three steps:

(1) the near ultraviolet and visible spectra of the ligands were recorded over the pH range of about 0.5 - 15.0,(representative spectra are reproduced in APPENDIX IV);

(2) the spectra were interpreted in conjunction with the data from the pH-potentiometric titrations of the ligands (representative titration data are found in APPENDIX III);

and (3) acid-dissociation constants (TABLE I) were calculated using the pH-potentiometric data.

2-PyOx. The spectra showed that only one species is predominant in the pH range -0.5 - 2.5. Since $\bar{p} = 2$ ($\bar{p}$ = average number of replaceable protons bound per molecule of ligand, as determined by potentiometric titration) in the pH range 2 - 2.5 (APPENDIX III), this species must be H_2L^+. The spectra and the values of $\bar{p}$ indicated that the species H_2L^+ lost a proton in the pH range 2.5 - 4.5. The pK_a value is similar to that of pK_a(NH) for oxine (TABLE I). Therefore, this equilibrium is associated with one of the nitrogen atoms of 2-PyOx. Since the pK_a(NH) values of pyridine and oxine are similar (5.15(62) and 5.02(63), respectively, in water at 20°C), the basic strengths of the two nitrogen atoms in 2-PyOx are likely similar also, and the following equilibrium is possible:

TABLE I

ACID-DISSOCIATION CONSTANTS

(50% (v/v) aqueous dioxane, 25.0°C, ionic strength = 0.1)

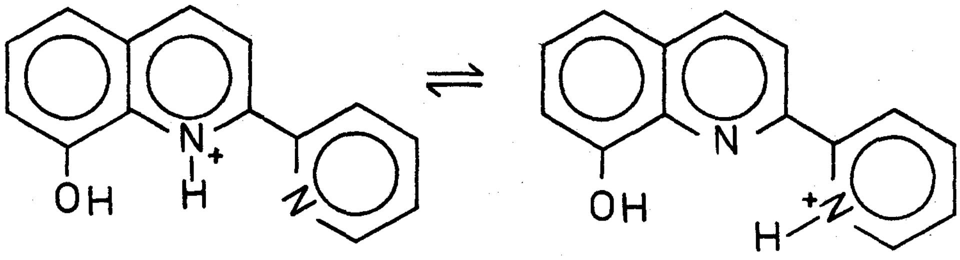
Compound	Equilibrium	$pK_a(NH)$	$pK_a(OH)$	Precision*
Oxine	$H_2L^+ \rightleftharpoons HL + H^+$	4.12	-	± 0.01
	$HL \rightleftharpoons L^- + H^+$	-	11.10	± 0.02
2-PyOx	$H_2L^+ \rightleftharpoons HL + H^+$	3.62 3.31**	-	± 0.01 ± 0.01
	$HL \rightleftharpoons L^- + H^+$	-	11.21 10.75**	± 0.02 ± 0.01
2-PhOx	$H_3L^+ \rightleftharpoons H_2L + H^+$	2.74	-	± 0.01
	$H_2L \rightleftharpoons HL^- + H^+$ (quinolinic)		10.28	± 0.02
	$HL^- \rightleftharpoons L^= + H^+$ (phenolic)		14.3***	± 0.3
2-HyOx	$H_2L \rightleftharpoons HL + H^+$	6.58	-	± 0.02
	$HL \rightleftharpoons L^- + H^+$	-	11.86	± 0.05

---

[*] Precision is reported as average deviation from mean (one titration).

[**] Refers to 60.0 C. These values were required for a temperature study of Ni(II) chelates of 2-PyOx.

[***] Calculated from ultraviolet spectra.



Hydrogen bonding between the two nitrogen atoms is also possible, but the difference between these situations is not important in the computation of metal-chelate formation constants.

In the pH range 10 - 12, the spectra and $\bar{p}$ values showed that HL dissociates to give L^-. Comparison of this pK_a with the pK_a values of oxine (TABLE I) shows that this equilibrium must be associated with the OH group of 2-PyOx.

The species H_3L^{2+} was not detected even at a pH of -0.5. These results are consistent with the acid-base reactions of 2,2'-bipyridyl and 1,10-phenanthroline in which the spatial arrangement of the nitrogen donors is similar to 2-PyOx. Beattie and Webster(64) have shown that for these two compounds, H_2L^{2+} is formed only in concentrated acid solutions.

Values of pK_a for oxines(30-32) show that 2-substituents usually have a smaller effect on the pK_a(OH) value than on the pK_a(NH) value. This is also true for 2-PyOx. The decrease in the pK_a(NH) value relative to oxine (∿ 0.5 log units) is expected, since the difference in pK_a(NH) values for pyridine and 2,2'-bipyridyl is 0.6 log units (pK_a(NH) for 2,2'-bipyridyl = 4.49 in water at 20°C(65)). The decrease in basicity can be attributed mainly to steric hindrance to solvation of the protonated nitrogen atom.

2-PhOx. The spectra and $\bar{p}$ values(APPENDIX III) showed that in the pH range -0.5 - 12 the species H_3L^+, H_2L and HL^- are obtained. From pH 13 to 15, the spectra indicated a further dissociation to L^{2-}. Since this equilibrium lies outside the pH range of the pH-potentiometric method, spectrophotometric techniques were used to estimate the pK_a value. The precision (± 0.3 log units) was poor for several reasons: (1) 2-PhOx decomposes slowly in strongly basic solutions, giving rise to an uncertainty in the absorbence measurements; (2) because of the limited solubility of sodium hydroxide (∿ 0.4 M) in 50% aqueous dioxane and the levelling effect of the solvent, complete dissociation of HL^- was not obtained; (3) the high sodium hydroxide concentrations required led to a variation in ionic strength; (4) in the medium used, the sodium hydroxide was not completely dissociated(48,49) and other ion-pair formation could be extensive; and (5) data from overlapping absorption bands were used in solving the simultaneous equations (see APPENDIX II for the calculations).

Since the pK_a(OH) of 2-(2'-hydroxyphenyl)-quinoline is 12.8(65) the pK_a(OH) values of 10.28 and 14.3 were assigned to the quinolinic and phenolic-OH groups respectively.

For 2-PhOx, both the pK_a(NH) and quinolinic pK_a(OH) decrease significantly relative to oxine. The decrease in pK_a(NH) could be attributed to steric hindrance to solvation of the protonated nitrogen atom, due to the 2-substituent(70) and intramolecular hydrogen bonding between the ring nitrogen atom and the 2-substituent-OH group. Evidence for hydrogen bonding is provided by Spectra 3 and 4 in Figure 4 which suggest that the broad absorption observed for 2-(2'-hydroxyphenyl)-8-

methoxyquinoline is due to hydrogen bonding of the phenolic proton with the ring nitrogen atom. The i.r. spectrum of 2-PhOx in the region 2500 cm^{-1} - 3500 cm^{-1} is similar to Spectrum 4 except for an additional strong absorption at 3200 cm^{-1} due to the quinolinic-OH group. Such a hydrogen bond would inhibit protonation of the nitrogen atom and increase its acidity. Similarly, intramolecular hydrogen bonding of the quinolinic-OH group to the nitrogen atom would also be inhibited, and a slight increase in the acidity of this group would also be expected*.

<u>2-HyOx</u>. The spectra of 2-HyOx vary in the pH range -0.5 - 1.5 and then remain essentially constant up to pH 5. Since $\bar{p} = 2$ (APPENDIX III) for the pH range 2.5 - 5, the alterations in the spectra below pH 2.5 likely correspond to the dissociation of H_3L^{2+}. This species is not important in the computation of metal-chelate formation constants and consequently the pK_a value was not calculated. In the pH ranges 6.0 - 7.5 and 11.0 - 12.5, the spectra and $\bar{p}$ values indicate the dissociation of H_2L^+ and HL, respectively. Since $pK_a = 7.92$ (in water at 25.0°C) for hydrazine(66), the pK_a value 6.58 was assigned to the hydrazino group.

The decrease in the pK_a(NH) value for the hydrazino group relative to the value in water is the result of solvent effects and possible intramolecular hydrogen bonding between the quinolinic-OH group and

---

*Molecular models show that hydrogen bonding between the 2-substituent-OH group and the nitrogen atom is more favorable than between the quinolinic-OH group and the nitrogen atom.

the hydrazino group*. As shown above the proton n.m.r. and i.r. spectra of 2-HyOx suggest the presence of a strong interaction between these two groups. This intramolecular hydrogen bonding would also be expected to increase the $pK_a(OH)$ value relative to the value for oxine. The greater acidity of the quinoline nitrogen compared to that in oxine is likely the result of steric and inductive effects attributed to the hydrazino group.

The standard deviation of the $pK_a(OH)$ value for 2-HyOx is larger than usual because of the slow decomposition of 2-HyOx in basic solutions. To minimize the effect of this decomposition the titrations were performed as rapidly as possible above pH 8.

The above explanations are necessarily qualitative and over-simplified. For example, no account of entropy differences has been made, yet such differences would certainly exist as a result of the different molecular structures and their different effects in structuring the solvent.

General Reactivity of Ligands

2-PhOx and 2-HyOx have been previously reported in the literature as potential antiseptic(54) and antitumoral drugs(58). Their chelating properties were not investigated. To gain some idea of the general reactivity of these ligands and of 2-PyOx towards metal ions, spot tests were performed with a large number of metal ions over a wide pH range. The spot tests were made at pH values of -0.3, 1.0, 2.6, 5.0, 7.5 and 10.0. The results are shown in TABLES II, III and IV respectively.

---

*Molecular models show that hydrogen bonding between these two groups could be important.

TABLE II

REACTIONS OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE WITH METAL IONS

Metal	pH=-0.3	pH=1.0	pH=2.6	pH=5.0	pH=7.5	pH=10
Al(III)	*		- **	+ ***	-	
Sc(III)			-	+	-	
Cr(III)			-	+	-	
Mn(II)			-	+	+	+
Fe(II)		-	-	+	+	+
Fe(III)	-	+	+	+	-	-
Co(II)			-	+	+	+
Ni(II)		-	-	+	+	+
Cu(II)		-	+	+	+	+
Zn(II)			-	+	+	+
Ga(III)		-	+	+	+	-
Sr(II)			-	-	-	
Zr(IV)			-	+	-	
Rh(III)			-	-	-	
Pd(II)	-	+	+	+	-	
Ag(I)			-	+	+	+
Cd(II)			-	+	+	+
In(III)			-	+	-	
La(III)			-	+	-	
Hg(II)	-	+	+	+	+	+
Pb(II)			-	+	+	+
Ce(III)			-	+	-	
Th(IV)			-	+	-	
U(VI)		-	+	+	+	+

---

*A blank space indicates that no test was made.

**A dash (-) indicates that a reaction was not observed.

***A plus sign (+) indicates that a reaction was observed (see Experimental).

TABLE III

REACTIONS OF 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE WITH METAL IONS

Metal	pH=-0.3	pH=1.0	pH=2.6	pH=5.0	pH=7.5	pH=10.0
Al(III)			-	-	-	+
Sc(III)			-	+	+	+
Cr(III)			-	-	-	-
Mn(II)			-	-	-	+
Fe(II)			-	+	+	+
Fe(III)	-	+	+	+	-	+
Co(II)			-	-	+	+
Ni(II)			-	-	-	+
Cu(II)		-	+	+	+	+
Zn(II)			-	-	+	+
Ga(II)		-	+	+	+	+
Sr(II)			-	-	-	-
Zr(IV)		-	+	+	-	-
Rh(III)		-	-	-	-	-
Pd(II)		-	+	-	-	-
Ag(I)			-	-	-	-
Cd(II)			-	-	-	+
In(III)			-	+	-	+
La(III)			-	+	+	+
Hg(II)			-	+	+	+
Pb(II)			-	-	-	+
Ce(III)			-	-	-	+
Th(IV)			-	+	-	+
U(VI)			-	+	-	-

TABLE IV

REACTIONS OF 2-HYDRAZINO-8-HYDROXYQUINOLINE WITH METAL IONS

Metal	pH=1.0	pH=2.6	pH=5.0	pH=7.5	pH=10.0
Al(III)		-	-	-	-
Sc(III)		-	-	-	-
Cr(III)		-	-	-	-
Mn(II)		-	-	-	+
Fe(II)		-	+	+	+
Fe(III)	-	+	+	+	+
Co(II)		-	+	+	+
Ni(II)		-	-	-	-
Cu(II)	-	+	+	+	+
Zn(II)		-	-	+	+
Ga(II)		-	-	-	-
Sr(II)		-	-	-	-
Zr(IV)		-	-	-	-
Rh(III)		-	-	-	-
Pd(II)		-	+	+	-
Ag(I)		-	reduces	reduces	-
Cd(II)		-	-	+	+
In(III)		-	-	-	
La(III)		-	-	-	-
Hg(II)		-	-	+	+
Pb(II)			-	-	+
Ce(III)		-	-	-	-
U(VI)		-	+	+	+

At pH 5, 2-PyOx reacts with 22 of the 24 metal ions tested and is unselective. The reactivity is greatly reduced at the lower pH values because of the competing donor-atom protonation reactions. At higher pH values, changes in reactivity can likely be accounted for in terms of more extensive metal-ion hydrolysis and differing complexation abilities of the various buffering agents used. At pH 5, 2-PhOx and particularly 2-HyOx are more selective than 2-PyOx, and this probably reflects less extensive metal-complex formations with these ligands because of the more basic nature of the ligand donor atoms. At higher pH values (e.g., pH 10), where the ligands are dissociated to a greater degree, the selectivity is poorer and is not significantly different from that of 2-PyOx. Thus, these tests suggest that the ligands will not likely prove to be selective analytical reagents.

An interesting aspect of TABLES II-IV is the observed reaction of Al(III) with 2-PyOx and 2-PhOx. The only 2-substituted oxines reported to react with Al(III) in aqueous solution are 2-methyloxine (2-MeOx)(29) and its derivatives(68). These reactions are observed only in the presence of a secondary ligand such as acetate, which prevents Al(III) hydrolysis as the pH of the solution is raised. The lack of reactivity of Al(III) with 2-MeOx in the absence of a secondary ligand has been explained in terms of steric interactions of the 2-substituent with coordinated water(24,26). Even in the presence of the acetate, however, oxines with large 2-substituent groups do not react with Al(III). For example, 2-phenyloxine (which is similar to 2-PyOx and 2-PhOx in both pK_a values and structure) does not react with Al(III) in aqueous solution(68)*. Since 2-PyOx and 2-PhOx do react

---

*This observation was confirmed in the present work.

(both in the presence and as shown below, in the absence of acetate), it would appear that the 2-substituents are coordinating. This implies that the ligands are either terdentate or are bidentate with the ring nitrogen and the substituent donor atoms serving as coordinating sites.

Another interesting feature is the reaction of 2-PhOx and 2-PyOx with U(VI). Oxine derivatives with 2-substituents of similar size (e.g., 2-phenyl and 2-(2'-thienyl)) do not form complexes with U(VI); hydrolysis of U(VI) occurs instead(69). Again, this behavior suggests that the 2-substituent donor group is an active donor and that the ligand is either terdentate or bidentate.

Differences in reactivity among the ligands as shown by TABLES II-IV likely reflects the relative ability of the 2-substituents to function as coordinating groups. For example, 2-PhOx shows greater reactivity with Sc(III), Ga(III), Zr(IV), In(III) and La(III).

This behavior is consistent with the observation that these ions are hard acids (class A metal ions) and that 2-PhOx is the hardest base (two oxygen donor atoms) of the three ligands. On the other hand, 2-PyOx and 2-HyOx are more reactive than 2-PhOx toward soft acids (class B metal ions) as Pd(II), Ag(I) and Cd(II).

Metal Chelates of 2-(2'-Hydroxyphenyl)-8-hydroxyquinoline

The reactions of 2-PhOx with Al(III), Co(II), Ni(II), Cu(II), Zn(II) and Cd(II), based on pH-potentiometric and spectrophotometric data, are discussed in this section and the reactions are compared to those of similar ligands. The metal chelate formation constants have been calculated and are listed in TABLE V. The formation constants for oxine chelates have also been determined and are listed in TABLE VI for comparison.

Reactions with Al(III). Three titration curves for the 2-PhOx/Al(III) system are shown in Figure 9. These were performed at ligand:metal-ion molar ratios of 1:1 (curve 3), 2:1 (curve 4) and 4.7:1 (curve 5). The titration curves for $HClO_4$ (curve 1) and the reagent (curve 2) are also given. The hydrolysis curves for the metal ions are shown in Figure 10.

In the titration of the 2-PhOx/Al(III) system (Figure 9), a buffer zone from about pH 3.7 - 5.0 was obtained, corresponding to the consumption of two moles of base per mole of metal ion, or the release of two protons per metal ion. Since the pH range of this buffer zone decreased with increasing molar ratios, the proton release must result from a complexation reaction such as

$$M^{2+} + HL \rightleftarrows ML^{+} + H^{+}$$

rather than from hydrolysis of Al(III) (Figure 10). Furthermore, since the titration curve for the 1:1 molar ratio (curve 3) displayed a similar buffer zone, the reaction in the above pH range must correspond to the formation of a 1:1 (terdentate) complex, i.e.,

$$Al^{3+} + H_2L \rightleftarrows AlL^{+} + 2H^{+}$$

Curves 4 and 5 show a second buffer zone, in the pH range 6-8, which also corresponds to a two proton release per Al(III) ion. Again, the pH range of this zone decreased with increasing molar ratio. This indicates the formation of AlL_2^-, i.e.,

$$AlL^{+} + H_2L \rightleftarrows AlL_2^{-} + 2H^{+}$$

The formation of AlL_2^- is also suggested by the shape of curves 4 and 5 beyond pH 8. This part of curve 4 can be superimposed on a strong acid

Figure 9. TITRATION CURVES OF ALUMINIUM(III) AND 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. (1) 5.205×10^{-4} moles $HClO_4$ in 110 ml 50% v/v aqueous dioxane (——), (2) as in 1 plus 5.08×10^{-5} moles 2PhOx (---··), (3) as in 1 plus 5.43×10^{-5} moles 2-PhOx plus 5.34×10^{-5} moles Al(III) (----), (4) as in 1 plus 1.08×10^{-4} moles 2-PhOx plus 5.34×10^{-5} moles Al(III) (-·-·), (5) as in 1 plus 2.52×10^{-4} moles 2-PhOx plus 5.34×10^{-5} moles Al(III) (····).

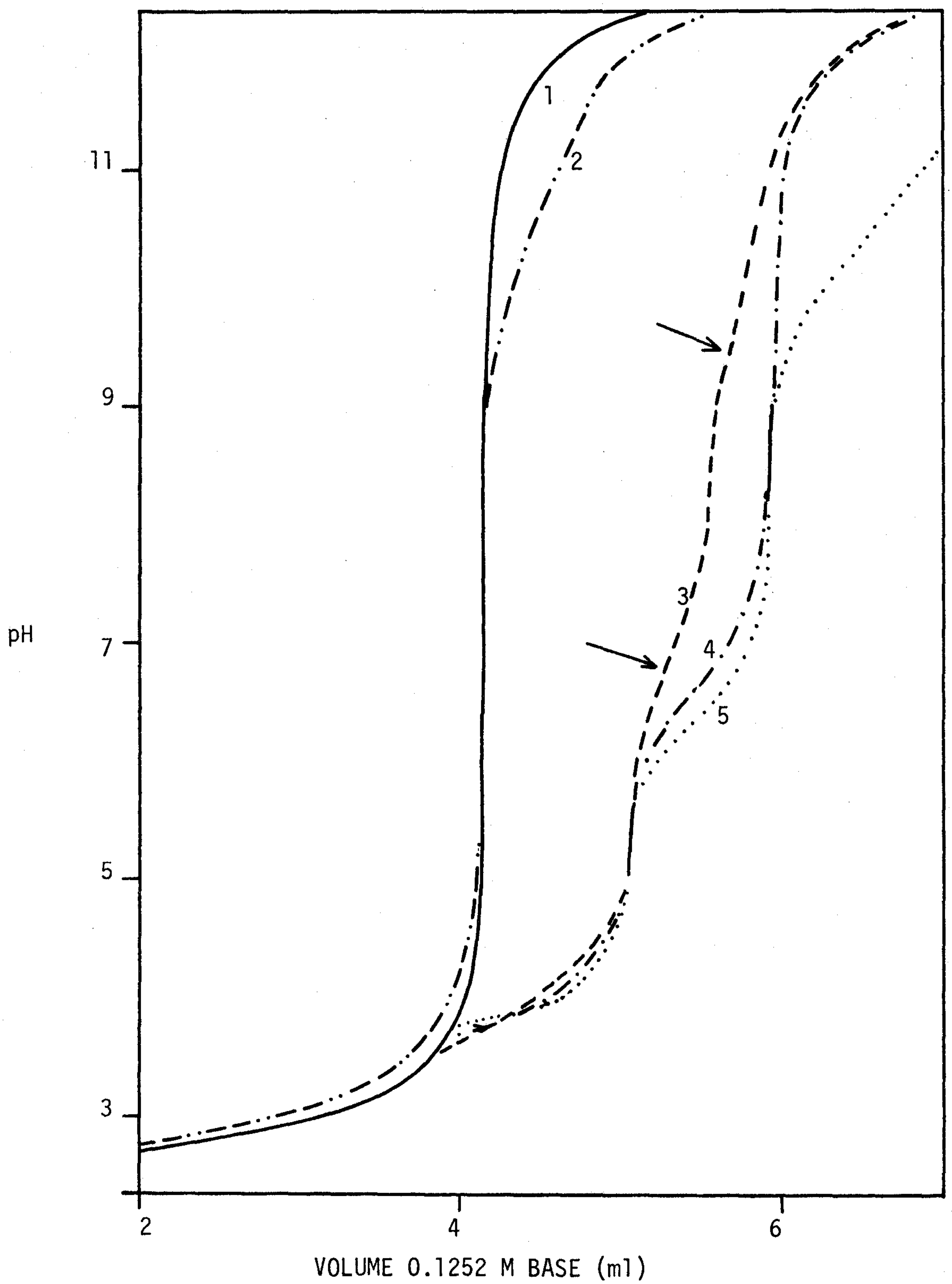
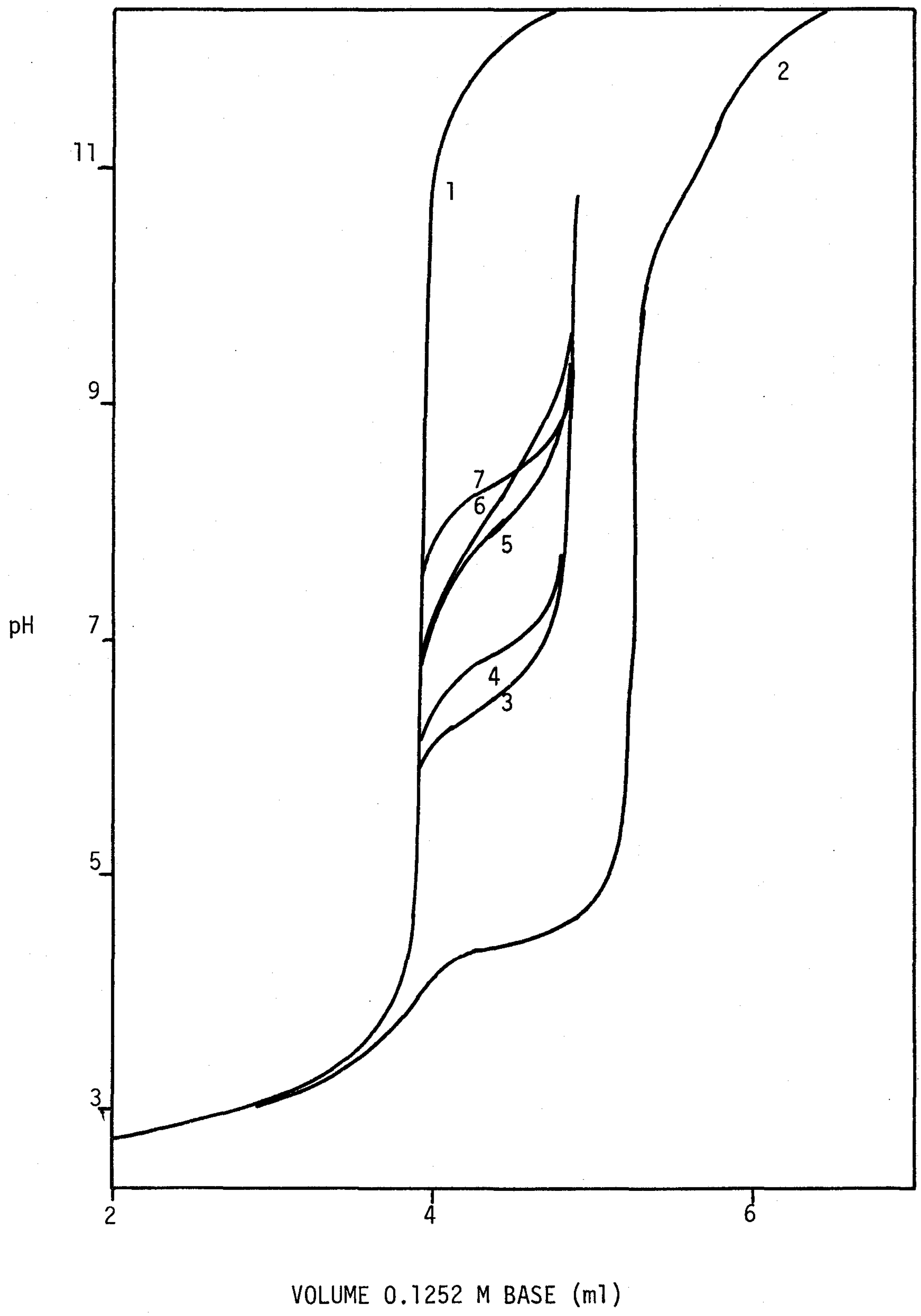

pH
3
5
7
9
11
2
4
6
VOLUME 0.1252 M BASE (ml)
1
2
3
4
5

Figure 10. METAL-ION HYDROLYSIS CURVES. (1) 4.93 x 10^{-4} moles $HClO_4$ in 110 ml 50% v/v aqueous dioxane, (2) as in 1 plus 5.34 x 10^{-5} moles Al(III), (3) as in 1 plus 5.40 x 10^{-5} moles Cu(II), (4) as in 1 plus 5.56 x 10^{-5} moles Zn(II), (5) as in 1 plus 5.23 x 10^{-5} moles Co(II), (6) as in 1 plus 5.46 x 10^{-5} moles Cd(II), (7) as in 1 plus 5.76 x 10^{-5} moles Ni(II).


11
9
7
5
3
pH
2
4
6
VOLUME 0.1252 M BASE (ml)
1
2
3
4
5
6
7

curve, indicating the absence of free 2-PhOx. The latter part of curve 5 corresponds exactly to the amount of free ligand calculated to be present on the basis of a 2:1 stoichiometry for the complex.

In the titration curve for 1:1 molar ratio (curve 3) two regions of proton release were observed above pH 5 (arrows), each corresponding to a one-proton release per metal ion. During the titration, a yellow-orange precipitate appeared at about pH 7 but redissolved above pH 9 to yield a clear yellow solution. These results suggest that AlLOH and $AlL(OH)_2^-$ form in these pH regions.

The following reactions are thus required to describe the titration curves:

$$Al^{3+} + H_2L \rightleftarrows AlL^+ + 2H^+$$

$$AlL^+ + H_2L \rightleftarrows AlL_2^- + 2H^+ \quad \text{for molar ratio 2:1}$$

$$AlL^+ + OH^- \rightleftarrows AlLOH \text{ (precipitate) for molar ratio 1:1}$$

$$AlLOH + OH^- \rightleftarrows AlL(OH)_2^-$$

Near pH 3.7 a small inflection can be seen for all three Al(III) titration curves. At this point the pH-meter readings decreased as the solution was allowed to stand. This behavior was observed throughout the first buffer zone. The presence of a solid phase is not responsible for these drifting readings, since no precipitate was observed. Furthermore, the observed drift is not due to simple Al(III) hydrolysis, since during the titration of Al(III) alone this feature was not observed.

The Al(III) hydrolysis curve (Figure 10) shows considerable deviation from the $HClO_4$ curve at pH 3.7, indicating partial hydrolysis of Al(III). The reactions which are involved in the hydrolysis of Al(III) have been studied extensively$^{(71-79)}$, but with widely divergent interpretations. However, most workers agree that polymeric hydroxo species predominate over mononuclear species, and that a characteristic of these species is their slowness in reaching equilibrium. Turner$^{(71)}$ has shown that the reaction between oxine and such polymeric Al(III) species is very slow. A similarly slow reaction with 2-PhOx would account for the drift in pH-meter readings with time. This condition means that the reaction between 2-PhOx and Al(III) (in polynuclear form) is mechanistically complex and that during titration it was not at equilibrium in the pH range corresponding to the first buffer zone. Interestingly, the same behavior was observed when Al(III) was titrated in the presence of 4-methyloxine which reacts with Al(III) in the same pH range as 2-PhOx.

Reactions with Co(II), Ni(II), Cu(II), Zn(II) and Cd(II). The titration curves for these metal ions are given in Figures 11 - 15. In each Figure are shown a $HClO_4$ curve (curve 1) and metal-complex titration curves for molar ratios of 1:1 (curve 2), 2:1 (curve 3) and about 3:1 (curve 4).

For each metal ion, a single buffer zone corresponding to two moles of base per mole of metal ion was observed below pH 8. The portion of the curves above pH 8 corresponded to titration of only that portion of the reagent in excess of a 1:1 molar ratio. For the titrations in a 1:1 molar ratio, the portion of the curves above pH 8 corresponded to the strong acid curve, although the Cu(II) and Ni(II) showed slight deviations. These results show that a 1:1 complex, ML, is formed with both ligand protons being liberated on chelation. The 1:1 stoichiometry of the Cu(II) and Ni(II) complexes was confirmed by Job's method of continuous variations. The Job

Figure 11. TITRATION CURVES OF COBALT(II) AND 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. (1) 5.16 x 10^{-4} moles $HClO_4$ in 110 ml 50% v/v aqueous dioxane (——), (2) as in 1 plus 5.15 x 10^{-5} moles 2-PhOx plus 5.23 x 10^{-5} moles Co(II) (----), (3) as in 1 plus 1.01 x 10^{-4} moles 2-PhOx plus 5.23 x 10^{-5} moles Co(II) (-·-·), (4) as in 1 plus 2.07 x 10^{-4} moles 2-PhOx plus 5.23 x 10^{-5} moles Co(II) (····).

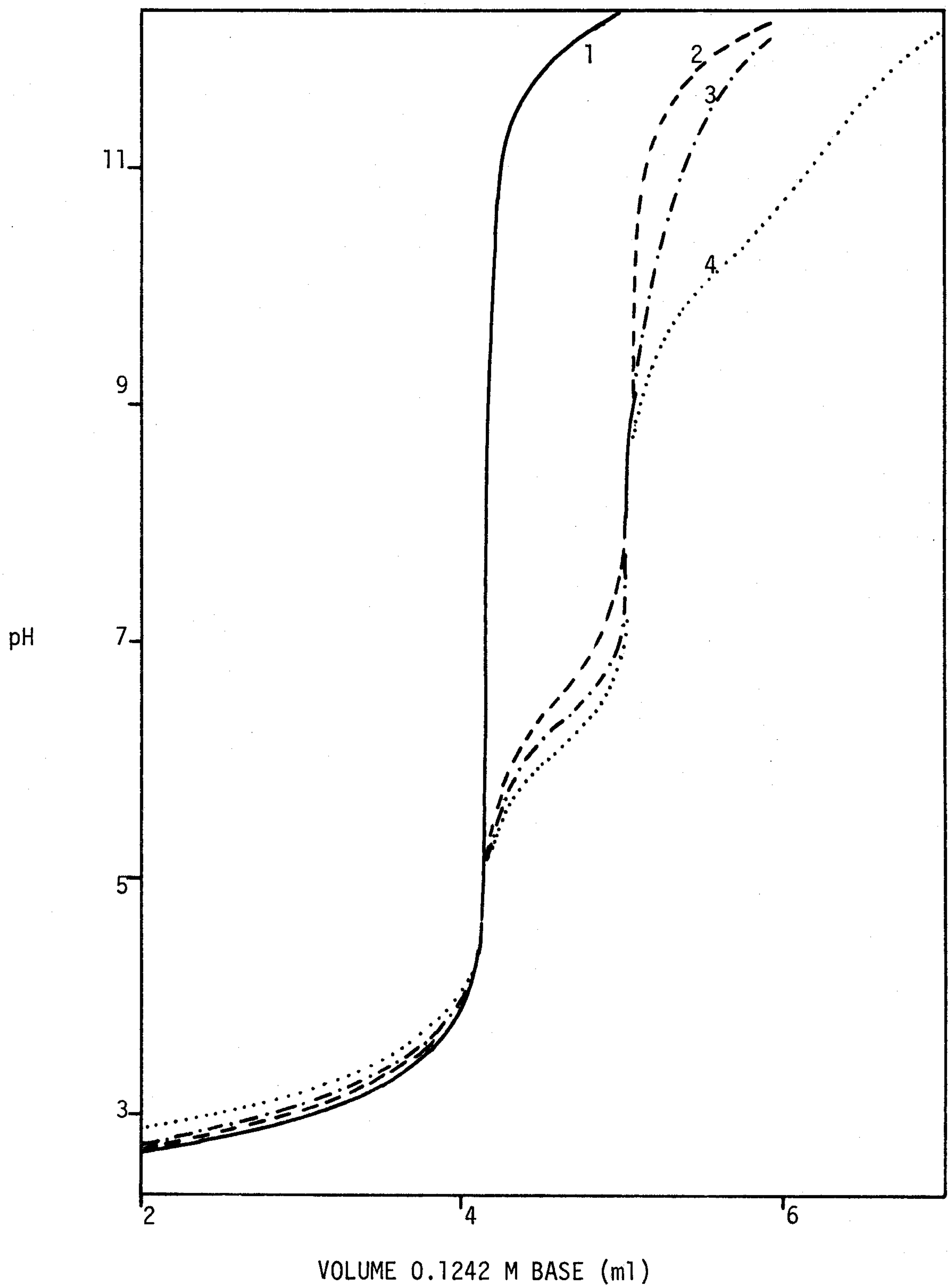
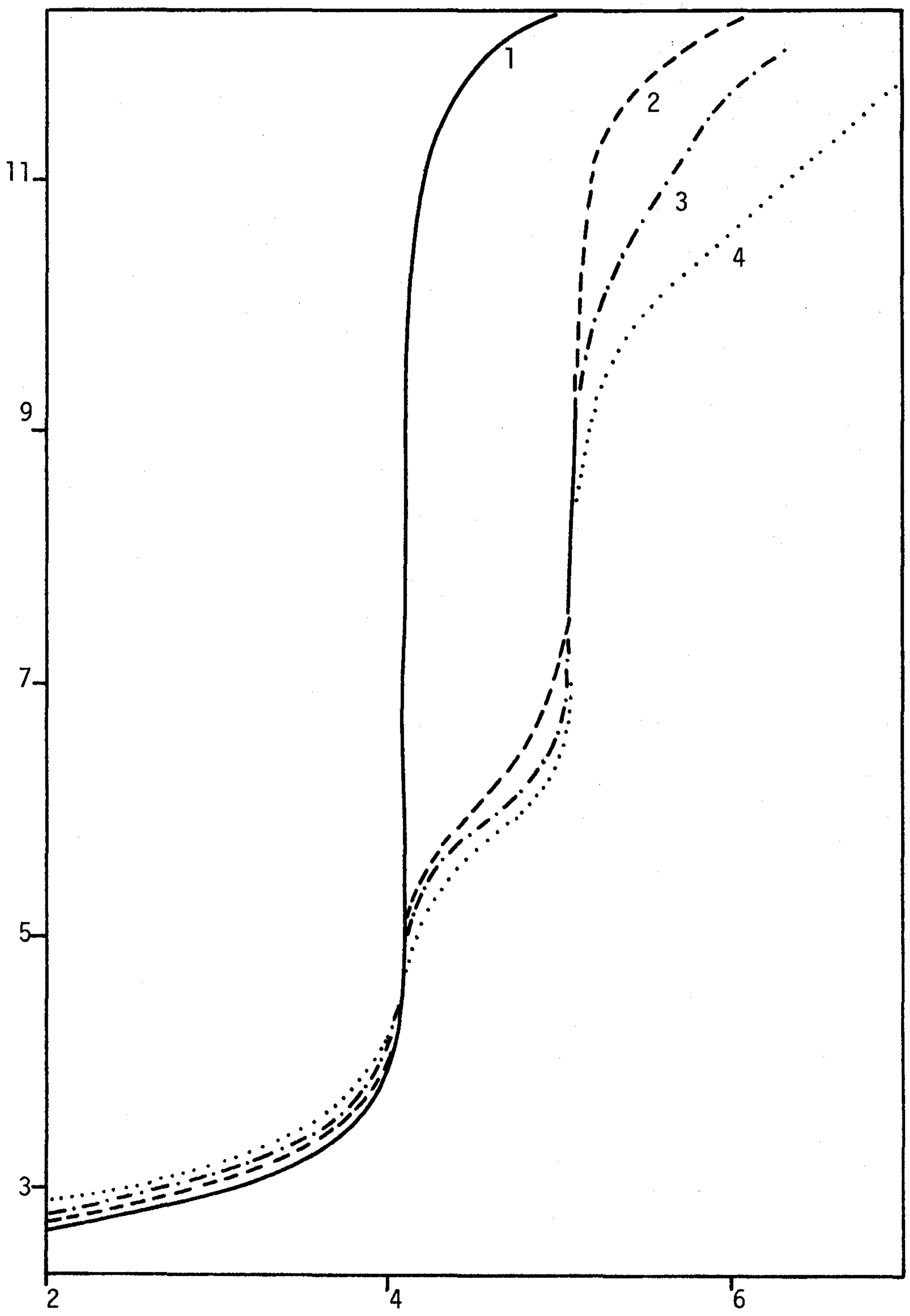

pH
11
9
7
5
3
2
4
6
1
2
3
4
VOLUME 0.1242 M BASE (ml)

Figure 12. TITRATION CURVES OF NICKEL(II) AND 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. (1) 5.16 x 10^{-4} moles $HClO_4$ in 110 ml 50% v/v aqueous dioxane (——), (2) as in 1 plus 5.82 x 10^{-5} moles 2-PhOx plus 5.76 x 10^{-5} moles Ni(II) (----), (3) as in 1 plus 1.15 x 10^{-4} moles 2-PhOx plus 5.76 x 10^{-5} moles Ni(II) (-·-·), (4) as in 1 plus 2.02 x 10^{-4} moles 2-PhOx plus 5.76 x 10^{-5} moles Ni(II) (····).



VOLUME 0.1242 M BASE (ml)

Figure 13. TITRATION CURVES OF COPPER(II) AND 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. (1) 5.16 x 10^{-4} moles $HClO_4$ in 110 ml 50% v/v aqueous dioxane (——), (2) as in 1 plus 5.64 x 10^{-5} moles 2-PhOx plus 5.40 x 10^{-5} moles Cu(II) (----), (3) as in 1 plus 1.07 x 10^{-4} moles 2-PhOx plus 5.40 x 10^{-5} moles Cu(II) (-·-.), (4) as in 1 plus 1.75 x 10^{-4} moles 2-PhOx plus 5.40 x 10^{-5} moles Cu(II) (····).

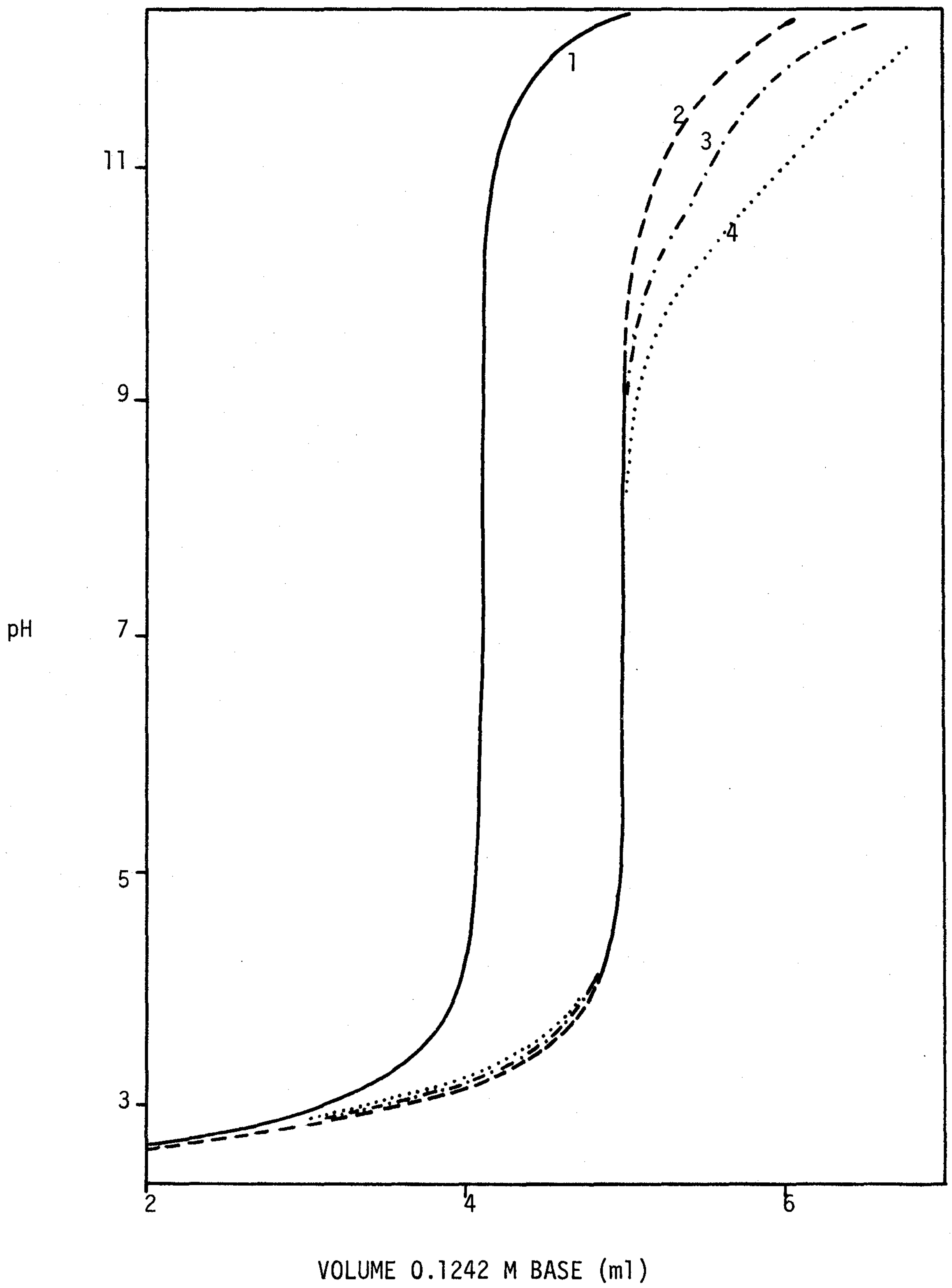

pH
3
5
7
9
11
2
4
6
1
2
3
4
VOLUME 0.1242 M BASE (ml)

Figure 14. TITRATION CURVES OF ZINC(II) AND 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. (1) 5.16 x 10^{-4} moles $HClO_4$ in 110 ml 50% v/v aqueous dioxane (——), (2) as in 1 plus 5.64 x 10^{-5} moles 2-PhOx plus 5.56 x 10^{-5} moles Zn(II) (----), (3) as in 1 plus 1.11 x 10^{-4} moles 2-PhOx plus 5.56 x 10^{-5} moles Zn(II) (-·-·), (4) as in 1 plus 2.12 x 10^{-4} moles 2-PhOx plus 5.56 x 10^{-5} moles Zn(II) (····).

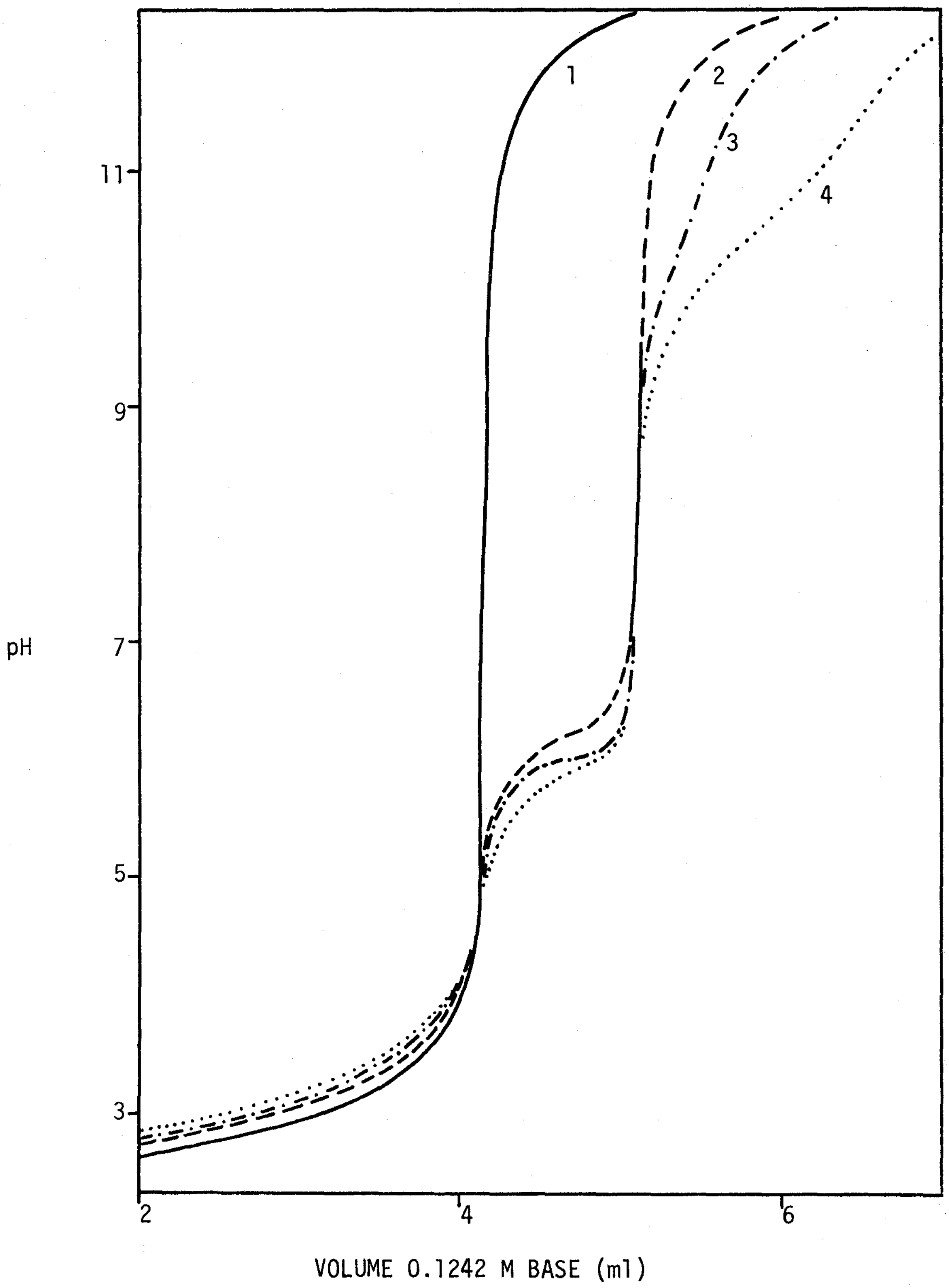
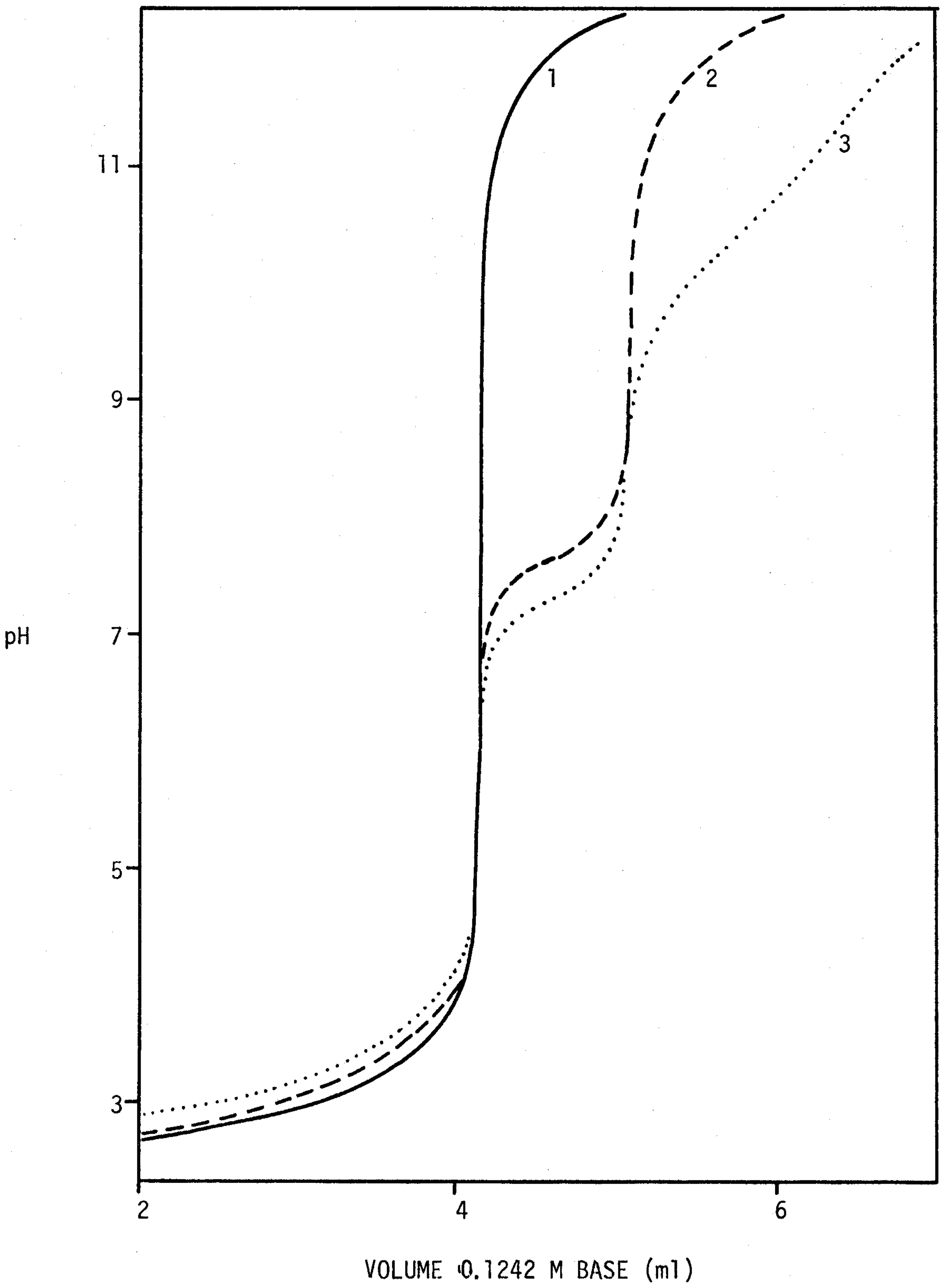

pH
3
5
7
9
11
1
2
3
4
2
4
6
VOLUME 0.1242 M BASE (ml)

Figure 15. TITRATION CURVES OF CADMIUM(II) AND 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. (1) 5.16 x 10^{-4} moles $HClO_4$ in 110 ml 50% v/v aqueous dioxane (——), (2) as in 1 plus 5.46 x 10^{-5} moles 2-PhOx plus 5.46 x 10^{-5} moles Cd(II) (----), (3) as in 1 plus 2.18 x 10^{-4} moles 2-PhOx plus 5.46 x 10^{-5} moles Cd(II) (····).


pH
11
9
7
5
3
2
4
6
VOLUME 0.1242 M BASE (ml)
1
2
3

plots were ideal ; spectrophotometric data for these plots were obtained at 290 nm and at pH 8 (i.e., beyond the pH range for the formation of ML). No potentiometric or spectrophotometric evidence was obtained to suggest formation of complexes of the type $ML_2^=$.

The following equation is required to describe the titration curves for all of the above metal ions:

$$M^{2+} + H_2L \rightleftarrows ML + 2H^+$$

Formation Constants of 2-PhOx Chelates. The formation curves ($\bar{n}$ versus pL) for the Al(III) chelate and the chelates of the other metal ions are shown in Figures 16 and 17, respectively. The formation constants are listed in TABLE V.

In the titration of the 2-PhOx/Cu(II) system, immediate precipitation occurred at molar ratios of 2:1 and greater because of the reduced solubility of ML. Consequently, the data used for formation constant computation was restricted to the 1:1 molar ratio titration. Precipitation of the Cd(II) and Zn(II) chelates (see Figure 17) occurred for all molar ratios, but sufficient data was obtained prior to precipitation to allow computation of the formation constants. Although the pH range (Figure 15) for the reaction between 2-PhOx and Cd(II) was near the pH range for Cd(II) hydrolysis (Figure 10), the precision of the formation constant calculated from data for titrations with different molar ratios was good. This agreement suggests that hydrolysis of Cd(II) was not a serious problem.

Figure 16 shows that the drift in pH values in the 2-PhOx/Al(III) system has caused distortion of the formation curve below $\bar{n}$ = 0.8. As a result, reliable values of K(AlL) could not be obtained. Values of log

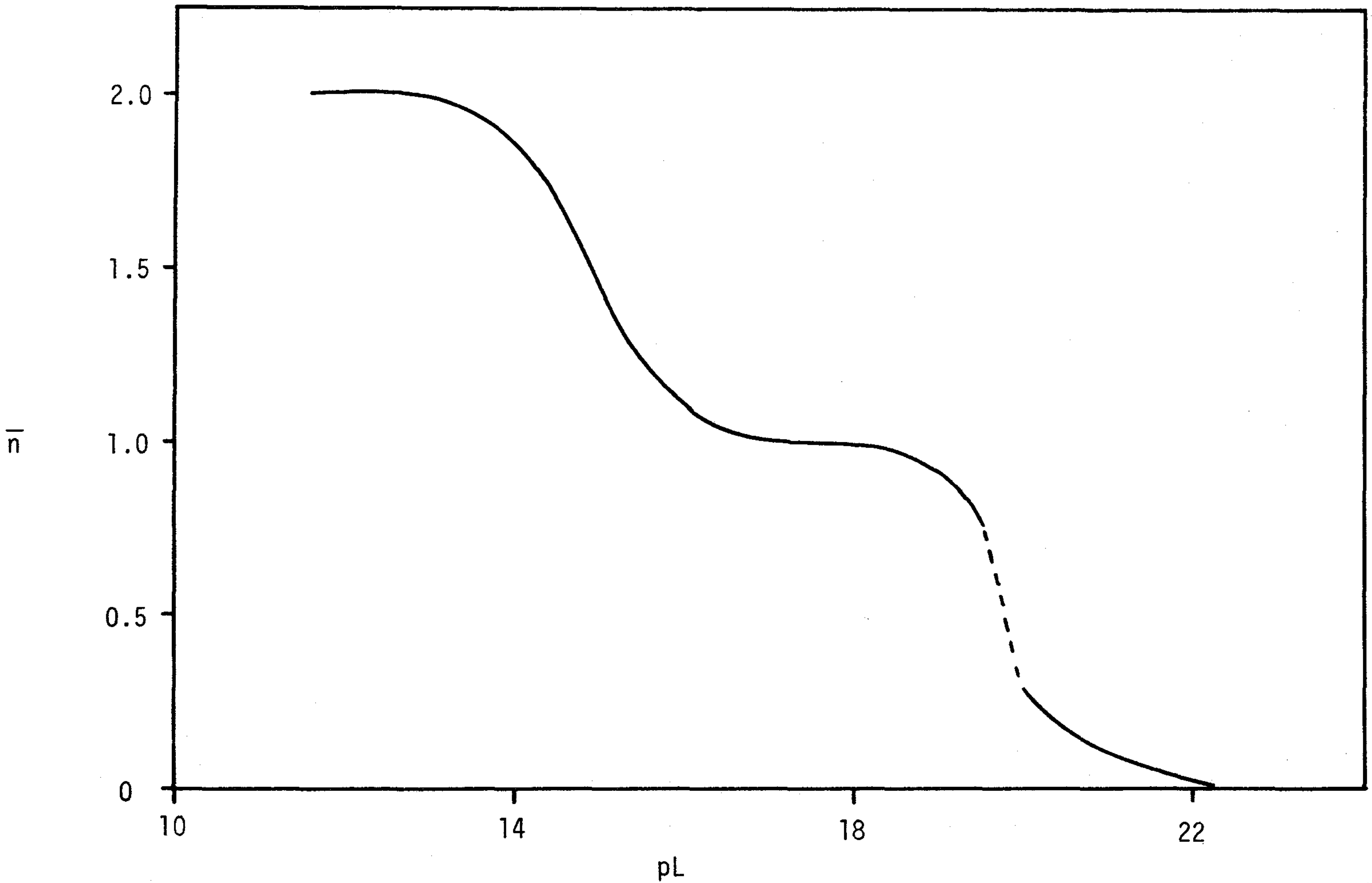


Figure 16. FORMATION CURVE FOR ALUMINUM(III) CHELATE OF 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. Broken line indicates drifting pH readings.

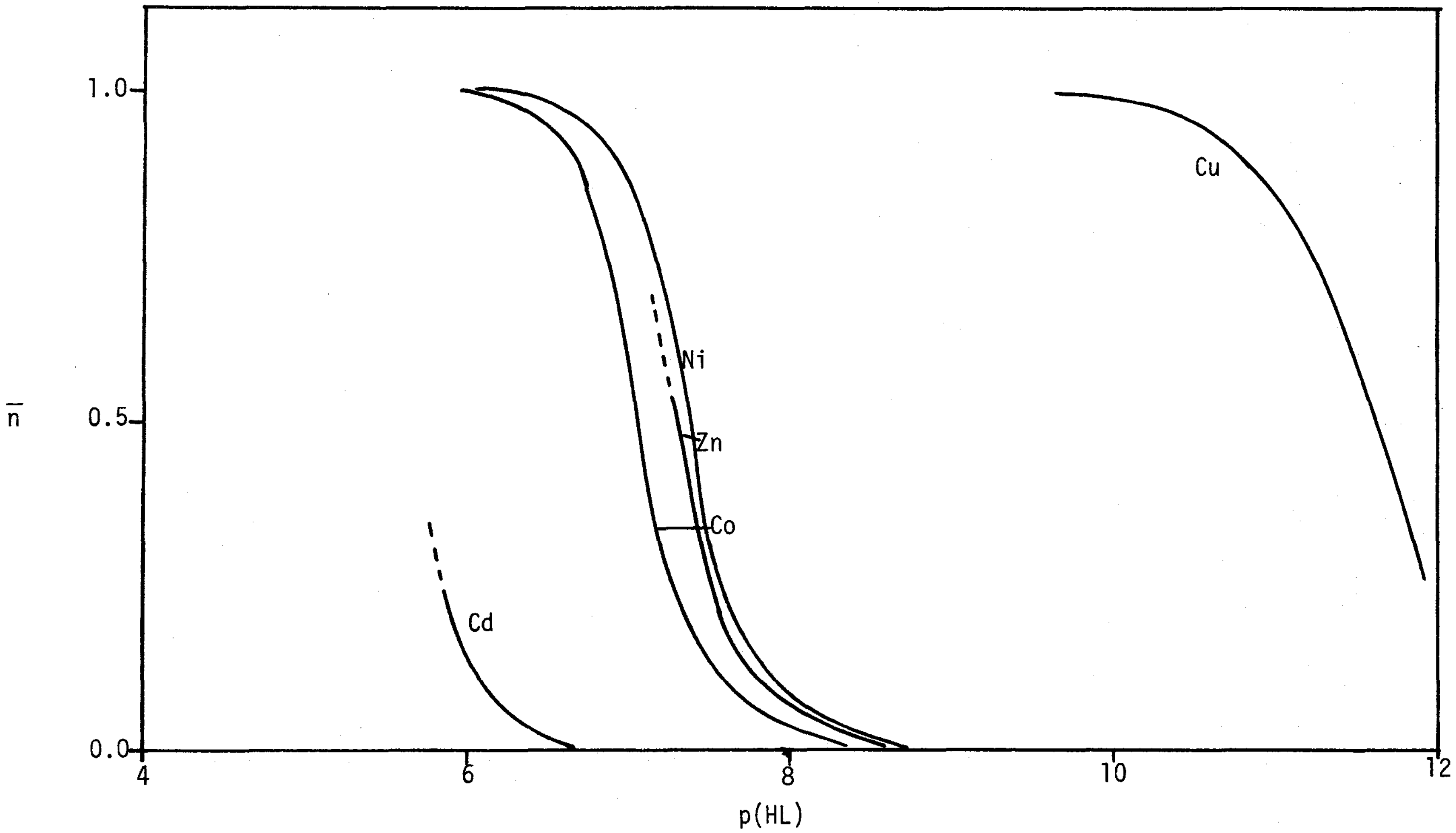


Figure 17. FORMATION CURVES FOR METAL CHELATES OF 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE. Broken line indicates precipitation.

TABLE V

FORMATION CONSTANTS OF 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE METAL CHELATES

(50% v/v aqueous dioxane, 25.0°C, ionic strength = 0.1)

Metal Ion	log K(ML)	log K(ML_2)	log β(ML_2)
Al(III)*	19.8 ± 0.5	14.9 ± 0.1	34.7
Co(II)	15.42 ± 0.02**	—	—
Ni(II)	16.00 ± 0.03	—	—
Cu(II)	23.60 ± 0.06	—	—
Zn(II)	15.82 ± 0.04	—	—
Cd(II)	12.57 ± 0.03	—	—

---

*Formation constants were approximated from the $\bar{n}$ curve.

**The precision is the average deviation from the mean of different titrations.

TABLE VI

FORMATION CONSTANTS OF 8-HYDROXYQUINOLINE METAL CHELATES (50% v/v aqueous dioxane, 25.0°C, ionic strength = 0.1)

Metal Ion	log K(ML)	log K(ML_2)	log K(ML_3)	log β(ML_2)	log β(ML_3)
Al(III)	11.98	11.02	10.74	23.00	33.74
Co(II)	9.66	8.05	—	17.71	—
Ni(II)	10.44	9.79	—	20.23	—
Cu(II)	13.39	12.24	—	25.63	—
Zn(II)	9.41	8.55	—	17.96	—
Cd(II)	8.22	6.95	—	15.17	—

K(AlL) and log K(AlL$_2$) were approximated, however, from Bjerrum's half-$\bar{n}$ method(80). For the oxine/Al(III) system, the successive formation constants decrease by 0.3 - 1.0 log units (TABLE VI), but for the 2-PhOx/Al(III) system, the difference between K(AlL) and K(AlL$_2$) is 4.8 log units (TABLE V). Molecular models suggest that ligand-ligand steric interactions should not be important in the 2-PhOx/Al(III) system. Consequently, the large difference is probably best explained in terms of electrostatic effects. For the 2-PhOx/Al(III) chelate, the charge of the reacting form of the metal ion is reduced by two units for each ligand bound, whereas for the oxine/Al(III) chelate, the charge is reduced by only one unit for each ligand bound.

As Adamson(81) has pointed out, it is incorrect to compare equilibrium constants when the units of these constants are different. Comparison of the K(ML) values of 2-PhOx with the K(ML) values of oxine can be made, however, since in each case the units are mole^{-1} liter. However, a direct comparison of the former with $\beta(ML_2)$ values of oxine is not valid since the units in the latter case are mole^{-2} liter^2. Conversion of the units to "unit mole fraction" is necessary(82). The resulting constants would then be dimensionless and a direct comparison of different systems could be made.

Thus, if the oxine and 2-PhOx complexes are compared using the displacement reaction

$$\mathrm{M(oxine)_2 + 2\text{-}PhOx \overset{K}{\rightleftarrows} M(2\text{-}PhOx) + 2\ oxine}$$

where the equilibrium constant, K, is given by

$$K = K(\mathrm{M(2\text{-}PhOx)})\ /\ \beta(\mathrm{M(oxine)_2}),$$

then K must be corrected for the disparity in the units of the formation constants, as described by Adamson(81). This correction is made as follows:

$$\Delta G^{\overset{1}{\circ}} = \Delta G^\circ + \Delta nRT \ln 55.5$$

$$= \Delta G^\circ + 2.36\ \Delta\ n$$

where $\Delta G^{\overset{1}{\circ}}$ = corrected free energy for K,

ΔG° = free energy for K calculated from uncorrected constants

and Δn = moles of products minus moles of reactants for the displaced reaction.

Since $\Delta n = 1$, 1.7 log units must be subtracted from log K. With this correction, the values of the formation constants in TABLES V and VI show that the 2:1 oxine chelates are more stable than the 1:1 2-PhOx chelates by 4-6 log units.

This difference in stability is not as large as would be expected (about 8-12 log units) if 2-PhOx functioned as a bidentate ligand and the 2-substituent served only as a sterically hindering group. In the latter event the K(ML) values for oxine would have been larger than the K(ML) values for 2-PhOx, due to steric hindrance(70). Thus, a comparison of formation constants suggests strongly that 2-PhOx functions as a terdentate ligand. This is consistant with the fact that on complexation two protons are released per molecule for ML formed.

Other 2-substituted oxines with potential donor atoms in the 2-substituent have been prepared and their reactions with metal ions studied(32,35,70). The majority of these ligands functions as bidentate ligands with the 2-substituent acting as sterically hindering groups. Consequently, the stability of the metal chelates formed by the majority of these ligands is reduced relative to that for the corresponding oxine chelates. 4-Amino-5-hydroxyacridine(70) and 2-aminomethyloxine(35) do form terdentate complexes. The stability of the acridine complexes, however, is lower(70) than that of the oxine complexes because of chelate ring strain due to the rigid nature of the ligand. On the other hand, the stability of the 2-aminomethyloxinates is greater than that of the corresponding oxinates(35). Here the enhanced stability arises from the ease with which the ligand forms terdentate complexes, i.e., flexibility of the side chain allows a favorable spatial orientation of the three donor atoms. 2-PhOx behaves like 2-aminomethyloxine in this regard.

Comparison of the Reactions of 2-PhOx and 2-MeOx with Al(III). 2-PhOx is the first 2-substituted oxine reported to react with Al(III) in the absence of secondary complexing agents. This reaction was easily demonstrated by potentiometric titration. The reactivity is a result of coordination by the potential donor group in the 2-substituent. This coordination compensates for any steric effects the 2-substituent may have.

Until recently, it was believed that 2-MeOx did not react with Al(III) (see INTRODUCTION), at least in aqueous solution. For example, Irving and Pettit(29) reported that in the potentiometric titration of the 2-MeOx/Al(III) system, only hydrolysis of Al(III) was observed. The lack of reactivity was explained mainly in terms of steric destabilization of the chelate. Other 2-substituted oxines have also been reported to be nonreactive with Al(III). Recently, however, the ligands 2-MeOx, 2,3-diMeOx, 2,5-diMeOx, 2,3,4-triMeOx and 2-Me-5-ClOx have been shown to react with Al(III) in alkaline aqueous solution in the presence of acetate to yield insoluble chelates(29,68). The difference in the reactivity of 2-substituted oxines with Al(III) in the presence and absence of OAc^- can be explained as follows:

As the pH is raised above the value of pK_a ($\approx$ 5) for HOAc the extent of formation of the OAc^-/Al(III) complexes, which prevent Al(III)-hydrolysis, remains constant since $[OAc^-]$ does not increase above this pH value. In the pH range 4.5-10.5, 2-MeOx*is present mainly in the form, HL. Consequently the extent of formation of $Al(2\text{-}MeOx)_3$ increases with pH since $[L^-]$ increases with pH. Eventually $[Al(2\text{-}MeOx)_3]$ is large enough for precipitation of the tris chelate to occur. In the absence of OAc^- only the precipitation of Al(III) hydrolysed species is observed. This occurs at a much lower pH range (4.5-5). Since Al(III) hydrolysis is favored at lower pH values and precipitation of $Al(2\text{-}MeOx)_3$ is favored at higher pH values, this implies that the ratio, $[Al(OH)_3]/[Al(2\text{-}MeOx)_3]$, changes with the pH of the solution. However, it can be shown from the equilibrium expressions for the formation of both $Al(OH)_3$ and $Al(2\text{-}MeOx)_3$ that this ratio should in fact be constant in the pH range 4.5-10.5. Thus if $Al(2\text{-}MeOx)_3$ is precipitated at a high pH then it should also be precipitated at a lower pH. The failure of 2-MeOx to react with Al(III) at lower pH values in the absence of acetate is likely due to the presence of polynuclear hydroxo-Al(III) species. As suggested previously for the reaction of 2-PhOx with Al(III), such polynuclear species appear to be kenetically favored. Consequently, precipitation of hydrolysed Al(III) species occurs.

Similar behavior has been reported in the case of Nb(V) and Ta(V)$^{(83)}$.

---

*A similar argument can also be used for any of the 2-methyloxines which were discussed above.

The apparent non-reactivity of these two ions with large complexing agents (metallochromic indicators) has been attributed to the formation of polynuclear hydroxo species. In the presence of "activators" or small complexing agents, however, Nb(V) and Ta(V) are complexed by the indicators. The "activators" function as catalysts by breaking the polymeric chains into more reactive, monomeric metal-ion species with which the large indicators can now react more rapidly.

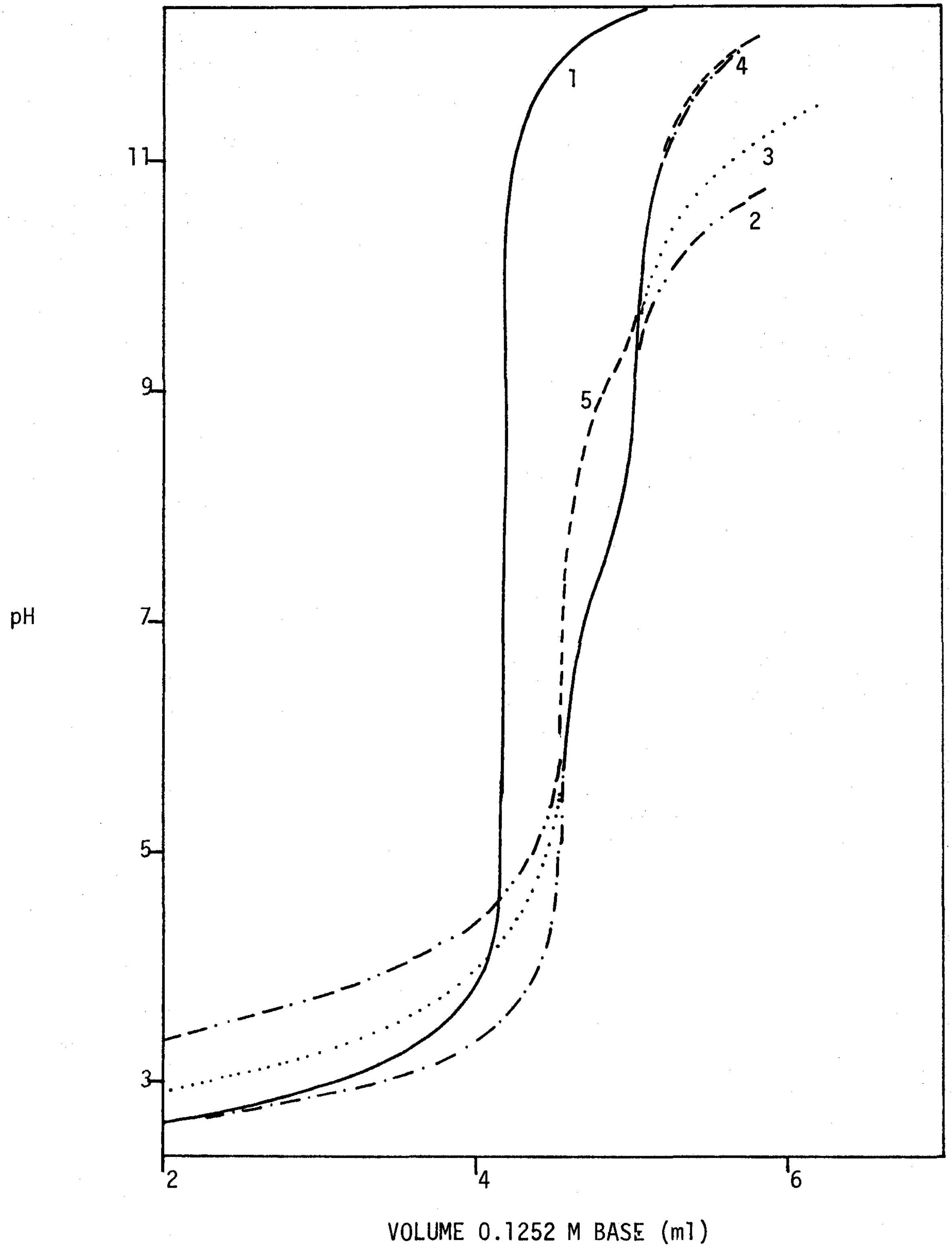
Metal Chelates of 2-(2'-Pyridyl)-8-hydroxyquinoline

The reactions of 2-PyOx towards Al(III), Co(II), Ni(II), Cu(II), Zn(II) and Cd(II) based on pH-potentiometric, spectrophotometric and polarographic techniques are discussed in this section. The metal-chelate formation constants have been calculated and are listed in TABLE VIII.

Reactions with Cu(II). Four titration curves for the 2-PyOx/Cu(II) system are shown in Figure 18. These were performed at ligand:metal-ion molar ratios of 9.3:1 (curve 2), 4.6:1 (curve 3), 2:1 (curve 4) and 0.9:1 (curve 5). The titration curve for $HClO_4$ (curve 1) is also given. The titration curves for 2-PyOx alone are given for three different concentrations in Figure 19.

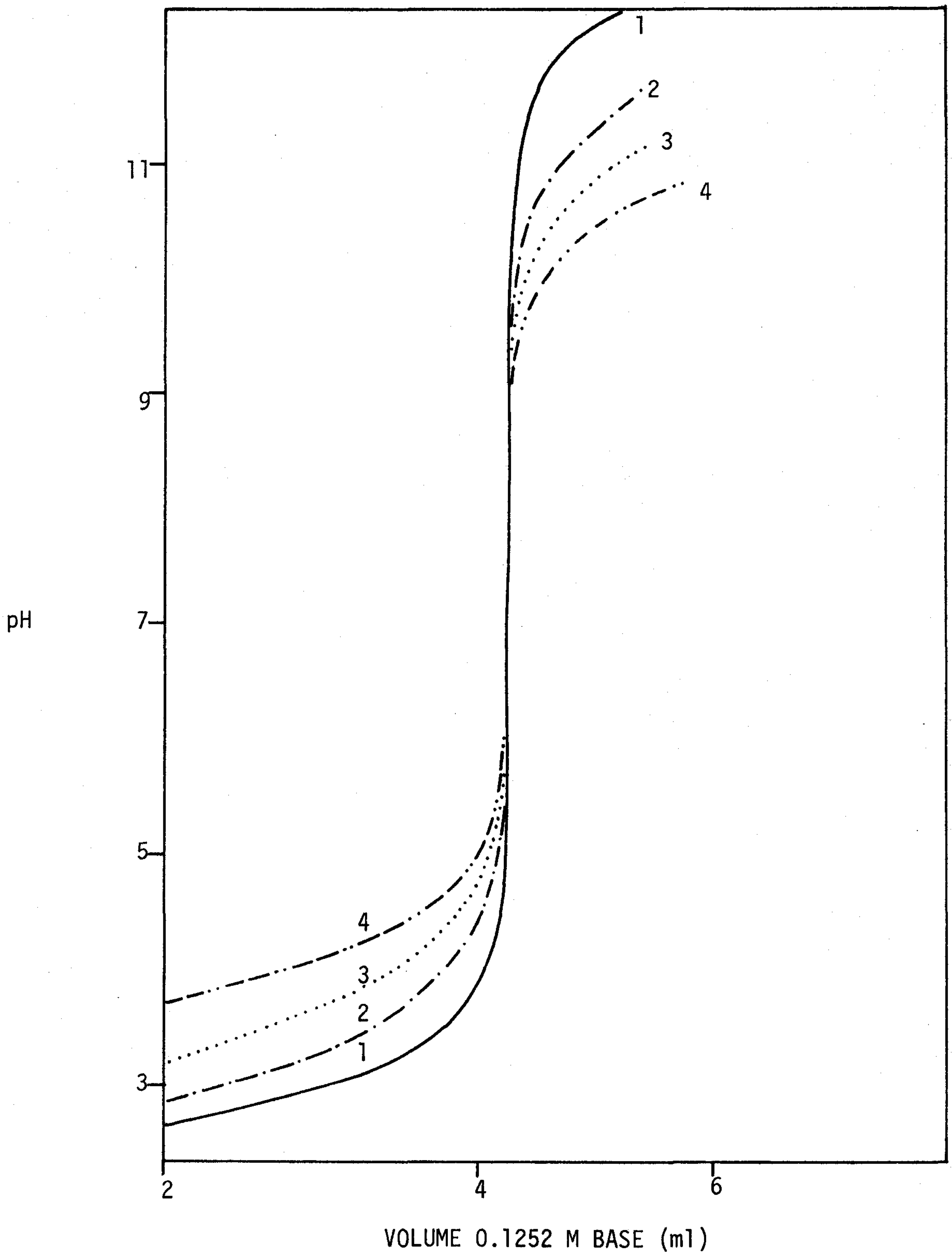
Curves 2, 3 and 4 for the 2-PyOx/Cu(II) system were found to be shifted from the strong acid curve in two steps. The first occurred from 4.15 to 4.60 ml and the second from 4.60 to 5.05 ml of base. Each step corresponds to the consumption of one mole of base per mole of metal ion. Above pH 10, curve 4 deviated only slightly from the strong acid curve (since no excess 2-PyOx was present), and curves 2 and 3 corresponded to the amount of free reagent calculated to be present on the basis of a 2:1

Figure 18. TITRATION CURVES OF COPPER(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. (1) 5.21 x 10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——), (2) as in 1 plus 5.04 x 10^{-4} moles 2-PyOx and 5.40 x 10^{-5} moles Cu(II) (--··), (3) as in 1 plus 2.50 x 10^{-4} moles 2-PyOx and 5.40 x 10^{-5} moles Cu(II) (····), (4) as in 1 plus 1.06 x 10^{-4} moles 2-PyOx plus 5.40 x 10^{-5} moles Cu(II) (-·-·), (5) as in 1 plus 5.04 x 10^{-5} moles 2-PyOx plus 5.40 x 10^{-5} moles Cu(II) (----).



pH
11
9
7
5
3
2
4
6
VOLUME 0.1252 M BASE (ml)
1
2
3
4
5

Figure 19. TITRATION CURVES OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. (1) 5.21×10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——), (2) as in 1 plus 5.04×10^{-4} moles 2-PyOx (---··), (3) as in 1 plus 2.50×10^{-4} moles 2-PyOx (····), (4) as in 1 plus 1.06×10^{-4} moles 2-PyOx (-·-·).


pH
11
9
7
5
3
2
4
6
VOLUME 0.1252 M BASE (ml)
1
2
3
4
4
3
2
1

stoichiometry for the complex. The above results suggest that the final complex formed in these titrations is CuL_2.

The buffer zone in the region 4.60 to 5.05 ml of base for the titration of a 1:1 molar ratio (curve 5) occurred at a significantly higher pH range than the corresponding buffer zones for the other titrations, suggesting the occurrence of hydrolysis and the formation of CuLOH.

The above results indicate that the reaction between 2-PyOx and Cu(II) in molar ratios $\geq$ 2:1 occurs with the stepwise formation of CuL^+ and CuL_2. However, in the pH range 6.5 - 7.5, curves 2, 3 and 4 can be superimposed upon one another. The formation constant

$$K = [H^+]\,[ML_2]\ /\ ([ML^+][HL])$$

shows that the titration curve for a 9.3:1 molar ratio should be $\sim$0.7 pH units below the titration curve for a 2:1 molar ratio, if CuL_2 is formed from CuL^+ in the above pH region. Thus, this pH region is not dependent on the free ligand concentration and CuL_2 must be formed from some species other than CuL^+.

Evidence for the formation of species other than CuL^+ and CuL_2 is also given by the formation curves for the 2-PyOx/Cu(II) system (Figure 20). These formation curves correspond to the titration curves in Figure 18. The shape of these formation curves is unusual since the same value of $\bar{n}$ is obtained at three different values of pL (negative logarithm of the free ligand concentration). This means that the equations which were used to describe the system are incorrect and that species other than CuL^+ and CuL_2 must be formed during titration. These formation curves are still useful, however, if they are interpreted in terms of the number and source of protons liberated on chelation, instead of the average

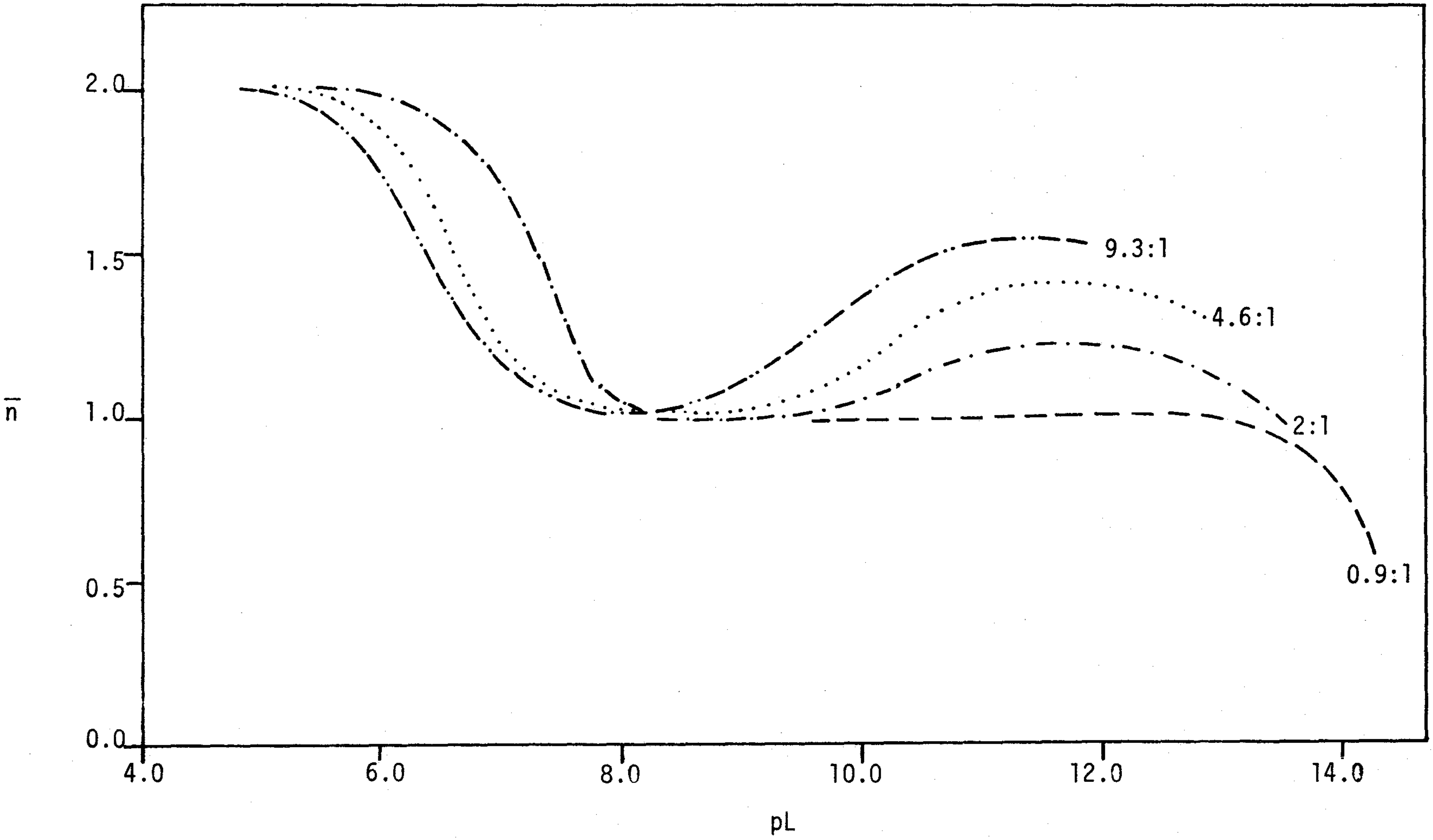


Figure 20. FORMATION CURVES FOR COPPER CHELATES OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. The ligand: metal-ion molar ratios are given to the right of the curves.

number of ligands bound per metal ion ($\bar{n}$), at a given pL value. The equations (APPENDIX II) which were used to calculate $\bar{n}$ for 2-PyOx/metal-ion systems were based on the ligand oxine. These equations were derived on the assumption that 2-PyOx is terdentate and that any protons bonded to the donor atoms are liberated on chelation. The magnitude of $\bar{n}$ at any point in the titration indicates both the number and the source of the liberated protons which is shown as follows. In acidic solutions two of the functional groups of 2-PyOx are protonated, but above pH 5 the only remaining acidic group is the quinolinic-OH group. Depending on the pH the following reactions can occur:

$$M^{2+} + H_2L^{+} \rightleftarrows ML^{+} + 2H^{+} \text{ (acidic solution)}$$

$$M^{2+} + HL \rightleftarrows ML^{+} + H^{+} \text{ (above pH 5)}$$

In acidic solution, a value of $\bar{n} = 1$ corresponds to the liberation of two protons per metal ion(from both the NH^+ and OH groups), but above pH 5 it corresponds to the liberation of only one proton which must come from the OH group.

The formation curve for a 1:1 molar ratio (Figure 20) attained a maximum constant value of one in acidic solution. Thus, both protons in 2-PyOx were liberated on chelation and CuL^+ was formed.

The remaining formation curves for the 2-PyOx/Cu(II) system are different. The $\bar{n}$ values increase to a maximum value greater than one and then decrease with decreasing pL values (i.e., with an increase in pH) to $\bar{n} = 1$ at pL $\approx$ 8 (pH $\approx$ 6). At this point (pH $\approx$ 6), only one OH proton was liberated per Cu(II) ion. Since $\bar{n} > 1$ below pH 6, CuL^+ has reacted with excess 2-PyOx. This reaction must involve the liberation of only NH^+

protons since the value of $\bar{n}$ at pH ≈ 6 shows that only one OH proton is liberated. Thus it is necessary to postulate the formation of $CuLHL^+$, where the proton in this species is associated with an oxygen donor in one of the 2-PyOx ligands. The reaction is:

$$CuL^+ + H_2L^+ \rightleftarrows CuLHL^+ + H^+ \text{ (below pH 6)}$$

If $CuLHL^+$ is the most stable species formed in acidic solutions (i.e., when the molar ratio 2-PyOx : Cu(II) is ≥ 2:1), $\bar{n}$ should attain a maximum value of 1.5 since the overall reaction involves the liberation of three protons. The formation curve for the 9.3:1 molar ratio does indeed have maximum $\bar{n}$ values ≈ 1.5 in acidic solution.

At pH ≈ 6, $CuLHL^+$ is in equilibrium with free 2-PyOx as follows:

$$CuLHL^+ \rightleftarrows CuL^+ + HL$$

Since this reaction does not involve the liberation of protons, the formation of $CuLHL^+$ (at pH ≈ 6) can not be detected when a glass electrode is used as a reaction probe. Thus $\bar{n}$ should decrease to unity from $\bar{n}$ ≈ 1.5, as observed when the pH is raised to 6.

It has been noted above that in the titrations with molar ratios ≥ 2:1 a buffer zone was obtained in the pH range 6.5 - 8.0. This buffer zone can be explained by the following reaction:

$$CuLHL^+ \rightleftarrows CuL_2 + H^+$$

The pH region in which this reaction occurs is not dependent on the free ligand concentration, provided the formation of $CuLHL^+$ is essentially complete prior to its acid-dissociation.

The 2-PyOx/Cu(II) system was also investigated by spectrophotometric techniques. Ultraviolet-visible absorption spectra of solutions containing 1:1 and 2:1 molar ratios of 2-PyOx:Cu(II) were recorded in 50% (v/v) aqueous dioxane over the pH range of 0 - 8. Spectra of 2-PyOx were also recorded for comparison (Figure 1, APPENDIX IV). The spectra of 2-PyOx showed no changes over the pH range 0 - 2. When Cu(II) was present in a 1:1 mole ratio, however, a new absorption appeared in the form of a shoulder at 310 nm (Figure 21). This absorption increased as the pH was increased to 4 and then remained constant in the pH range 4 - 8. This band is similar to the band at 295 nm in the spectrum of the 2-PyOx anion (L^-) (Figure 1, APPENDIX IV). The spectra indicated that no free 2-PyOx was present in the pH range 4 - 8. These results are consistent with the titration and formation curves for a 1:1 molar ratio, and suggest that fully deprotonated 2-PyOx, in the form of CuL^+, is responsible for the absorption at 310 nm.

Figure 22 shows the spectra for a 2:1 molar ratio. As above, the absorption at 310 nm appeared at pH 2 and then increased up to pH 4. There was no alteration in the spectra from pH 4 to 5. However, the spectra in this pH range differed from the corresponding spectra for a 1:1 molar ratio by the presence of maxima at 290 and 265 nm. At pH 8, these maxima were absent and the spectrum was similar to that obtained for a 1:1 molar ratio in the pH range 4 - 8. The molar absorptivities of the 310 nm maximum were equal in each case. Thus, at pH 8, the spectrum for a 2:1 molar ratio corresponds to the formation of CuL_2. The absorption maxima at 290 and 265 nm at pH 5 are indicative of the formation of $CuLHL^+$ or of an equimolar mixture of CuL^+ and HL since they are similar to those present in the spectrum of the neutral form of 2-PyOx (APPENDIX IV). The spectral data favor the formation

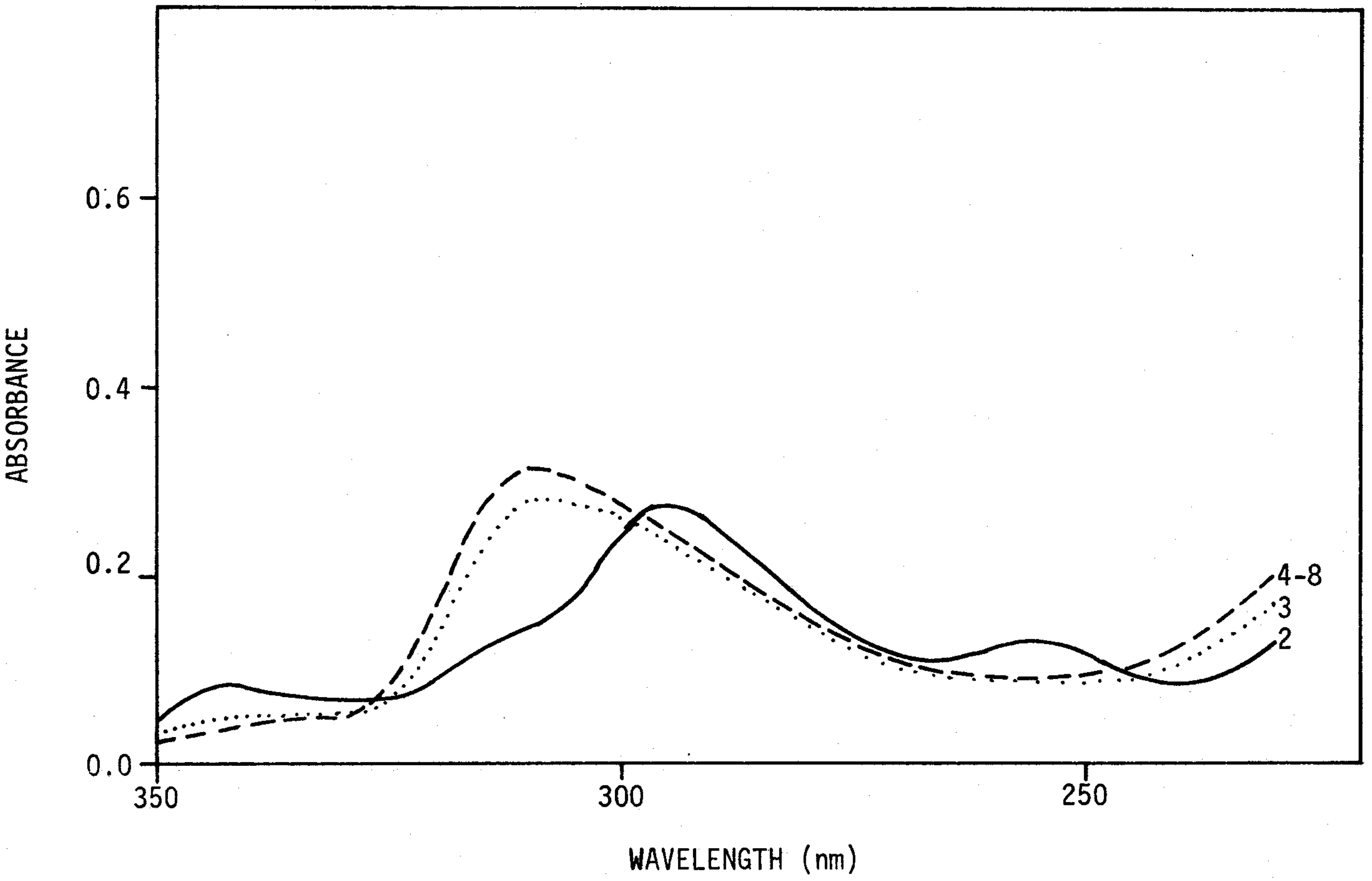


Figure 21. SPECTRA OF COPPER(II) PERCHLORATE AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE IN A 1:1 MOLAR RATIO (1.0 x 10^{-5} M). The pH of the solutions are given to the right of the curves.

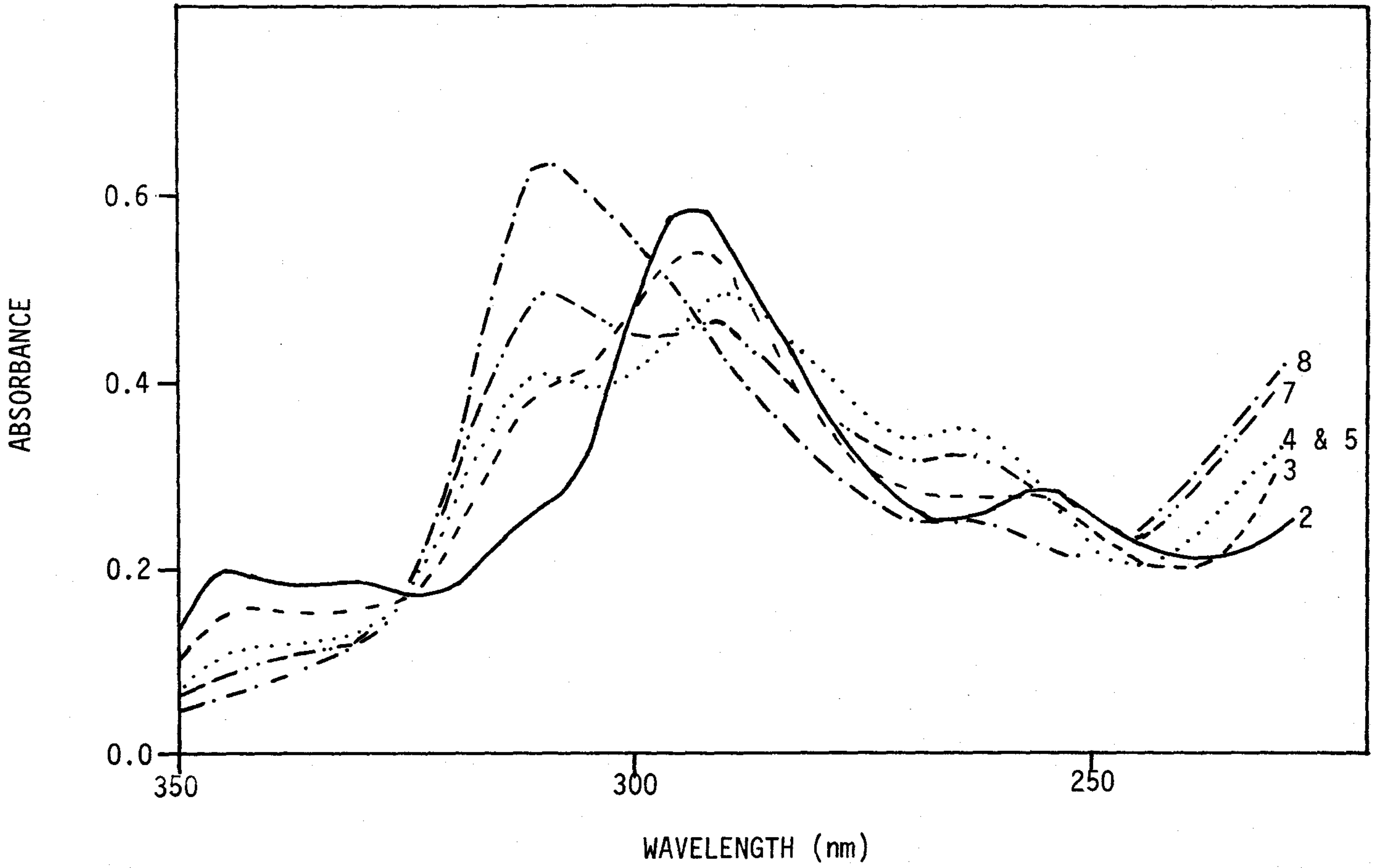


Figure 22. SPECTRA OF COPPER(II) PERCHLORATE AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE IN A 2:1 MOLAR RATIO. The pH of the solutions are given to the right of the curves. The concentrations are 2.0 x 10^{-5} M and 1.0 x 10^{-5} M for ligand and metal ion respectively.

of $CuLHL^+$, however, for the following reasons:

(1) the spectra of 2-PyOx changed significantly in the pH range 4 - 5 but the spectra for the 2:1 system were constant in this pH range;

(2) Figure 23 shows that at pH 5 the spectrum obtained by adding the absorbances of two solutions, one of which contained CuL^+ and the other an equimolar quantity of 2-PyOx, was not the same as the spectrum of a solution which contained a 2:1 molar ratio of the reacting species;

and (3) a mole-ratio plot in which the 2-PyOx concentration was constant shows that both a 2:1 and 1:1 species can exist at pH 5 and that the 2:1 species was formed only when the 2-PyOx concentration was in excess of that required for a 1:1 molar ratio (Figure 24).

The potentiometric data from five titrations, in which various initial metal-ion and 2-PyOx concentrations and various metal-ion:ligand molar ratios were used, were refined with the program SCOGS. An attempt was made to fit other species to the data but only CuL^+, $CuLHL^+$ and CuL_2 were found to be consistent with the data. Thus, the following sequence summarizes the reactions which occur between 2-PyOx and Cu(II):

$$Cu^{2+} + HL \rightleftarrows CuL^+ + H^+ \qquad pH\ 3.5$$

$$CuL^+ + HL \rightleftarrows CuLHL^+ \qquad pH\ 3 - 5$$

$$CuLHL^+ \rightleftarrows CuL_2 + H^+ \qquad pH\ 7 - 9$$

Reactions with Ni(II). Four titration curves for the 2-PyOx/Ni(II) system are shown in Figure 25. These were performed at ligand:metal-ion molar ratios of 5:1 (curve 2), 2:1 (curve 3) and 1:1 (curves 4 and 5). The titration curves for the systems containing 5:1 and 2:1 molar ratios were similar to the corresponding titration curves for the Cu(II) system

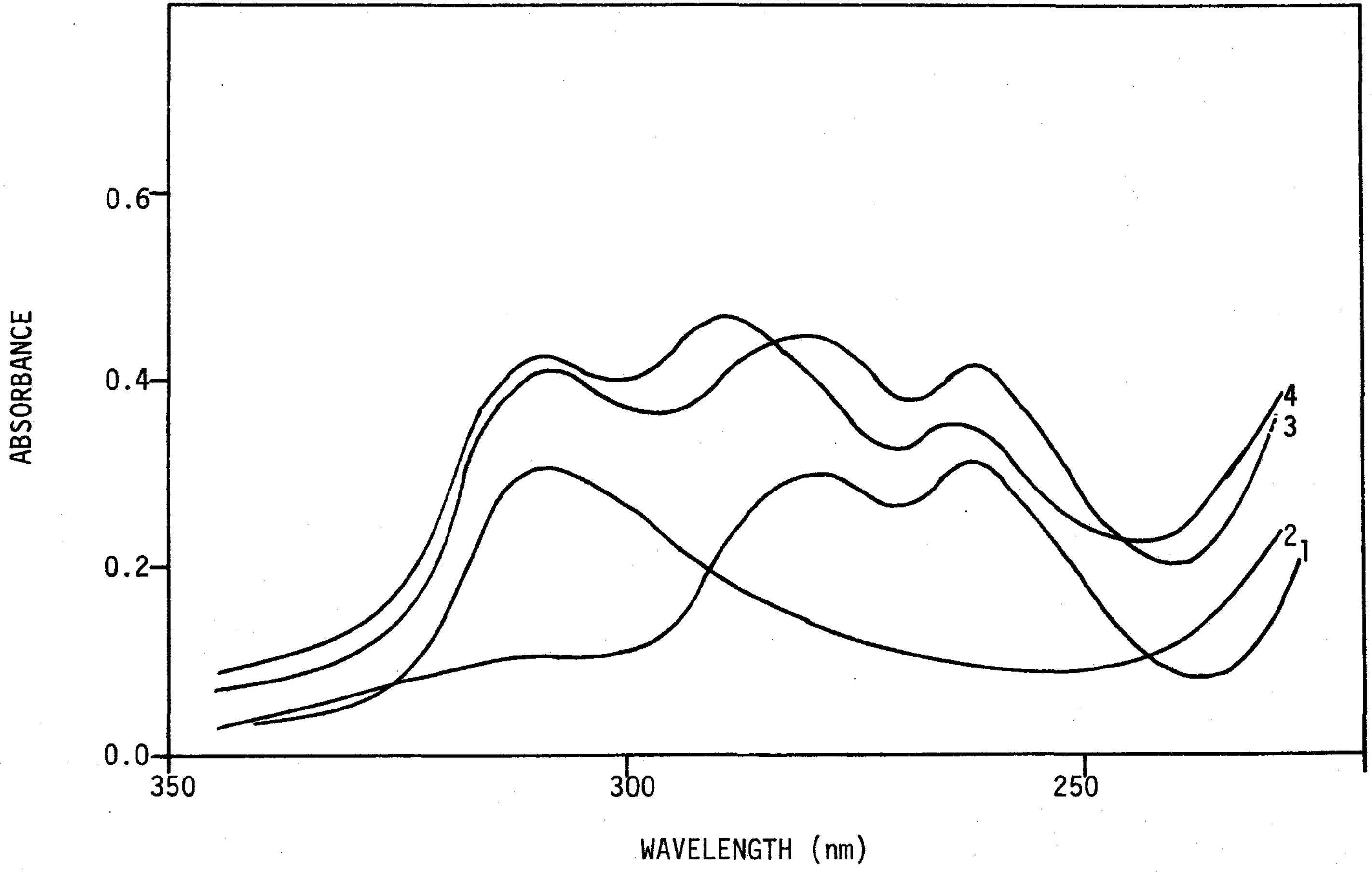


Figure 23. SPECTRA OF COPPER(II) PERCHLORATE AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. (1) [2-PyOx] = 1.0×10^{-5}, (2) as in 1 and [Cu^{2+}] = 1.0×10^{-5}. (3) sum of absorbances from 1 and 2, (4) [2-PyOx] = 2.0×10^{-5} M and [Cu^{2+}] = 1.0×10^{-5} M.

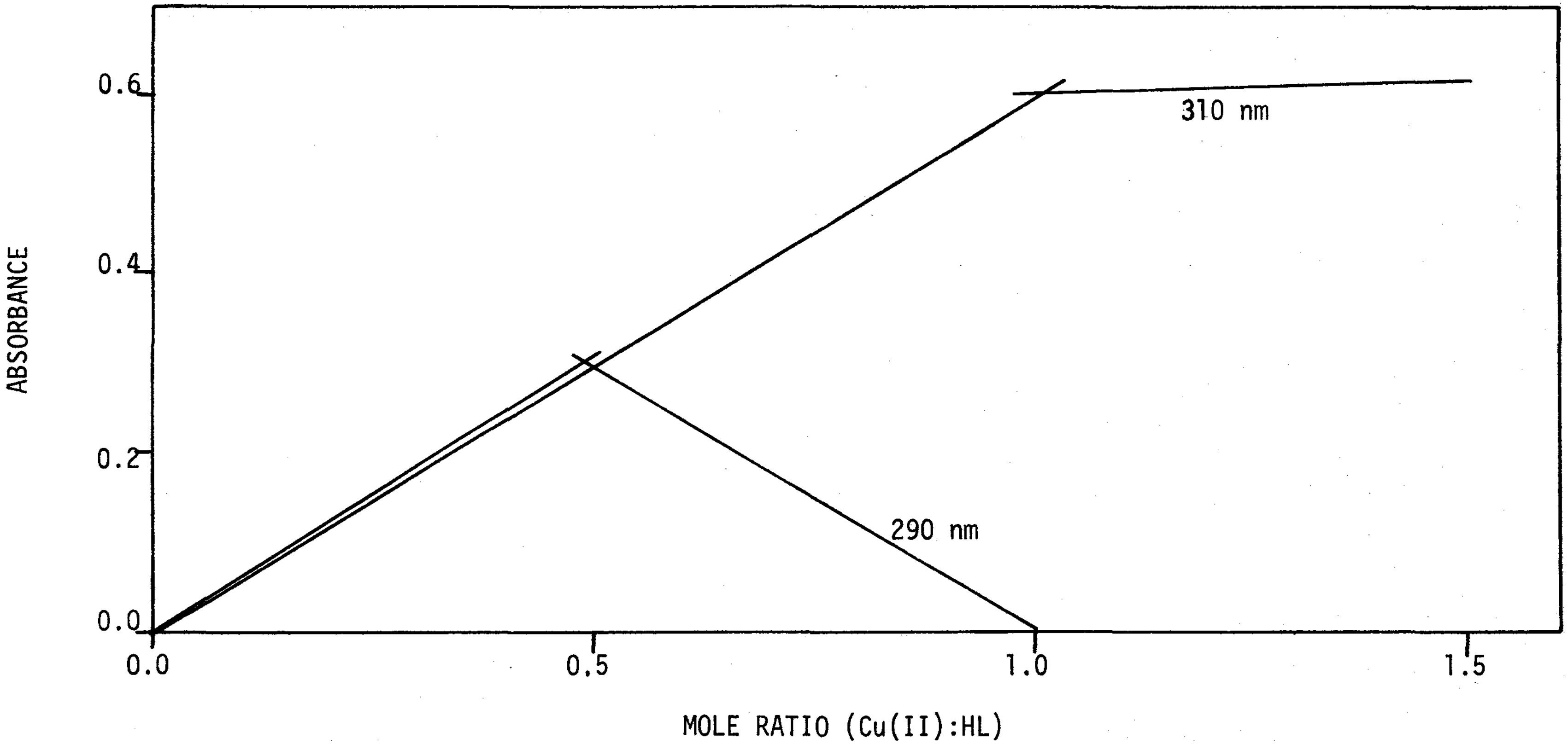
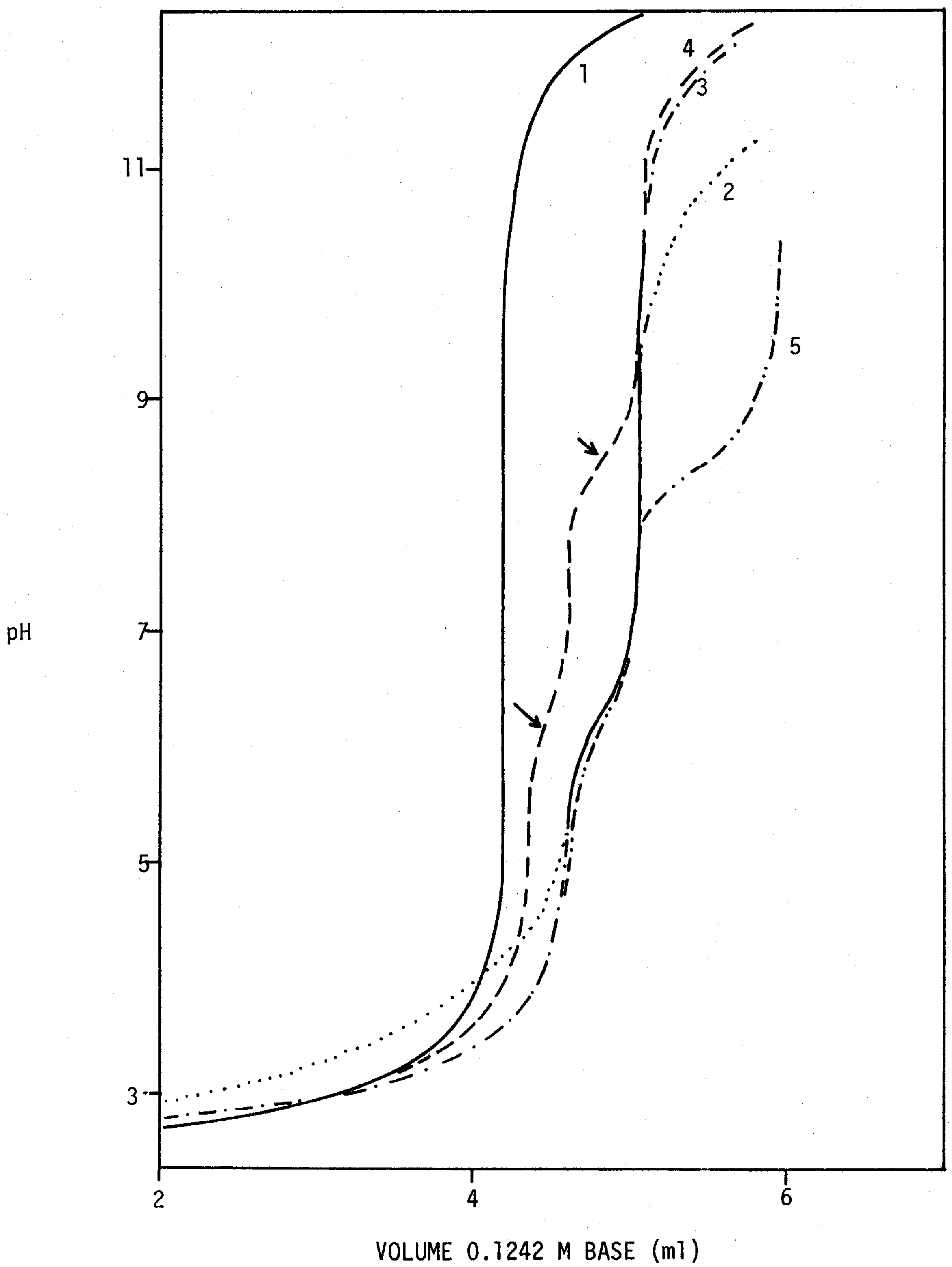


Figure 24. MOLE RATIO PLOTS FOR COPPER(II) CHELATES OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE AT pH 5. [2-PyOx] = 2.0 x 10^{-5}.

Figure 25. TITRATION CURVES OF NICKEL(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. (1) 5.16 x 10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——), (2) as in 1 plus 2.81 x 10^{-4} moles 2-PyOx and 5.76 x 10^{-5} moles Ni(II) (····), (3) as in 1 plus 1.18 x 10^{-4} moles 2-PyOx and 5.76 x 10^{-5} moles Ni(II) (-·-·), (4) as in 1 plus 5.54 x 10^{-5} moles 2-PyOx and 5.76 x 10^{-5} moles Ni(II) (----), (5) as in 1 plus 10 ml 50% (v/v) aqueous dioxane and 1.15 x 10^{-4} moles 2-PyOx and 1.15 x 10^{-4} moles Ni(II) (--··).


pH
11
9
7
5
3
2
4
6
VOLUME 0.1242 M BASE (ml)
1
2
3
4
5

(Figure 18, curves 3 and 4). These titration curves were shifted from that for $HClO_4$ by two buffer zones, each of which corresponded to the consumption of one mole of base per mole of metal ion. The pH range (5.5 - 7.0) of the second buffer zone was not altered by changes in the 2-PyOx concentration. The above results suggest that, as for the Cu(II) system, the Ni(II) system involves the species NiL^+, $NiHLH^+$ and NiL_2.

The titration curves for the 1:1 molar ratios were different from the corresponding titration curve for Cu(II) (Figure 18, curve 5). For Ni(II), the section of curve 4 from 4.15 to 4.60 ml base had two buffer zones (arrows). Each of these buffer zones corresponded to the consumption of one-half mole of base per mole of metal ion. Curve 5 shows that the pH region of the second buffer zone (pH 5.5 - 7.0) is similar to the second buffer zone of curves 2 and 3. This suggests that $NiLHL^+$ was also formed in the titrations of the 1:1 molar-ratio system. To explain the two buffer zones it is necessary to postulate that at pH 5, $NiLHL^+$ is considerably more stable than NiL^+ under the experimental conditions used for the potentiometric titrations. With this postulate the reactions occurring at pH 5 in a solution containing a 1:1 molar ratio can be written as follows:

$$Ni^{2+} + 2HL \rightleftarrows NiLHL^+ + H^+ \quad \text{pH } 3 - 5$$

$$NiLHL^+ \rightleftarrows NiL_2 + H^+ \quad \text{pH } 5 - 7$$

From these equations it is clear that one-half of the initial amount of Ni(II) would be uncomplexed and that one-half mole of base per mole of metal ion would be consumed for each reaction. In this manner, the titration curves can be explained.

The formation curves for the 2-PyOx/Ni(II) system are given in

Figure 26. These curves correspond to the titration curves shown in Figure 25. The formation curves for the 2:1 and 5:1 molar-ratio systems were similar to the corresponding formation curves in Figure 20 for the Cu(II) system. However, the 1:1 molar ratio curve was different in that $\bar{n}$ values of 0.5 rather than 1.0 were obtained over a wide range of pL values. For example, at pH 5 (pL ≈ 9.5), $\bar{n} = 0.5$ showing that only one-half of the available quinolinic-OH protons were liberated at this pH value. In more acidic solutions (e.g., pH ≈ 3, pL ≈ 12.5) $\bar{n}$ increases above 0.5 which, as was shown for Cu(II), is indicative of the formation of a protonated complex. To explain this formation curve it is necessary to postulate that the stability of $NiLHL^+$ is substantially greater than that of NiL^+ and that, as a result of this stability, $NiLHL^+$ predominates over NiL^+. This result is consistent with the above explanation of the 1:1 molar ratio titration curves.

The ultraviolet-visible spectra of the 2-PyOx/Ni(II) system in 1:1 and 2:1 molar ratios were recorded in 50% (v/v) aqueous dioxane over the pH range of 0 - 8. Figure 27 shows the main features of the spectra obtained for the 1:1 system. The spectra of 2-PyOx are shown in APPENDIX IV. The spectra for the 2:1 system were similar to the spectra for the 1:1 system and are not given. This similarity indicates that the same reactions occurred in both systems. The above spectra are also similar to the spectra of the 2-PyOx/Cu(II)system in a 2:1 molar ratio (Figure 21). Figure 28 shows mole-ratio plots for the 2-PyOx/Ni(II) system at pH 5. Unlike the 2-PyOx/Cu(II)system the graphs at each wavelength show that a 2:1 complex was formed. This 2:1 species was dominant even for small 2-PyOx:Ni(II)

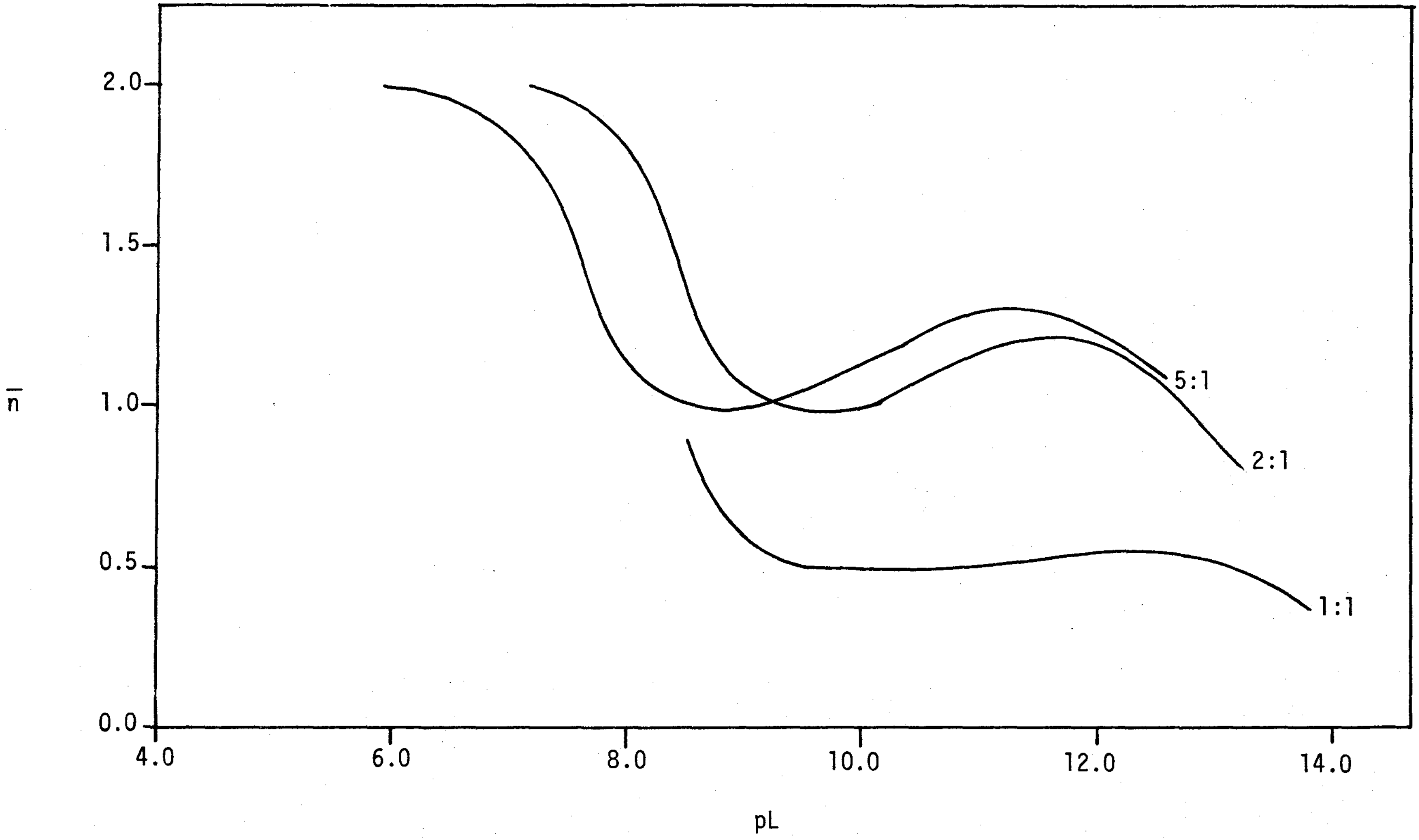


Figure 26. FORMATION CURVES FOR NICKEL(II) CHELATES OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. The ligand:metal molar ratios are given to the right of the curves.

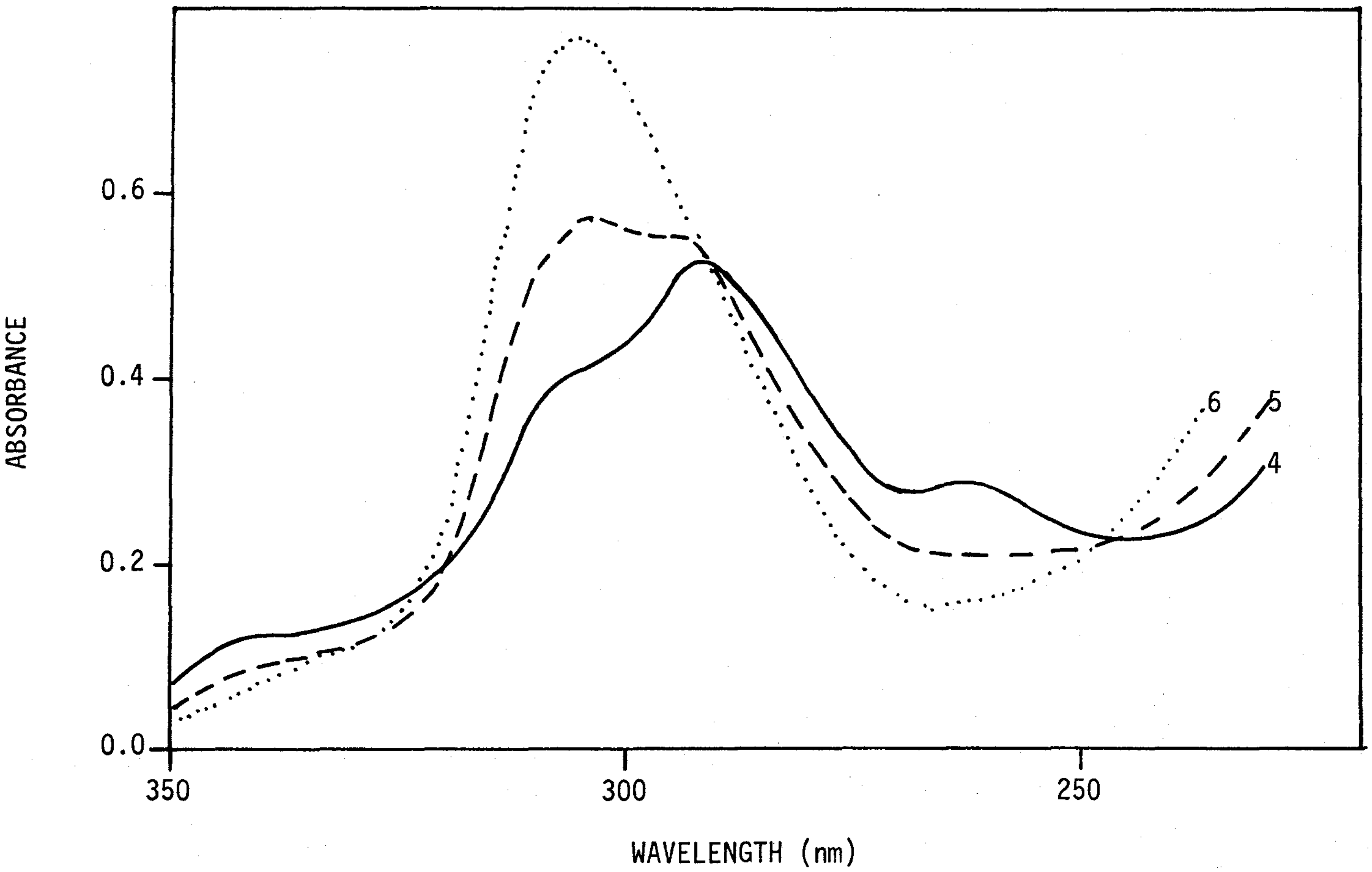


Figure 27. SPECTRA OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE AND NICKEL(II) PERCHLORATE IN A 1:1 MOLAR RATIO. The pH is given to the right side of the curves, [2-PyOx] = 2 x 10^{-5} M.

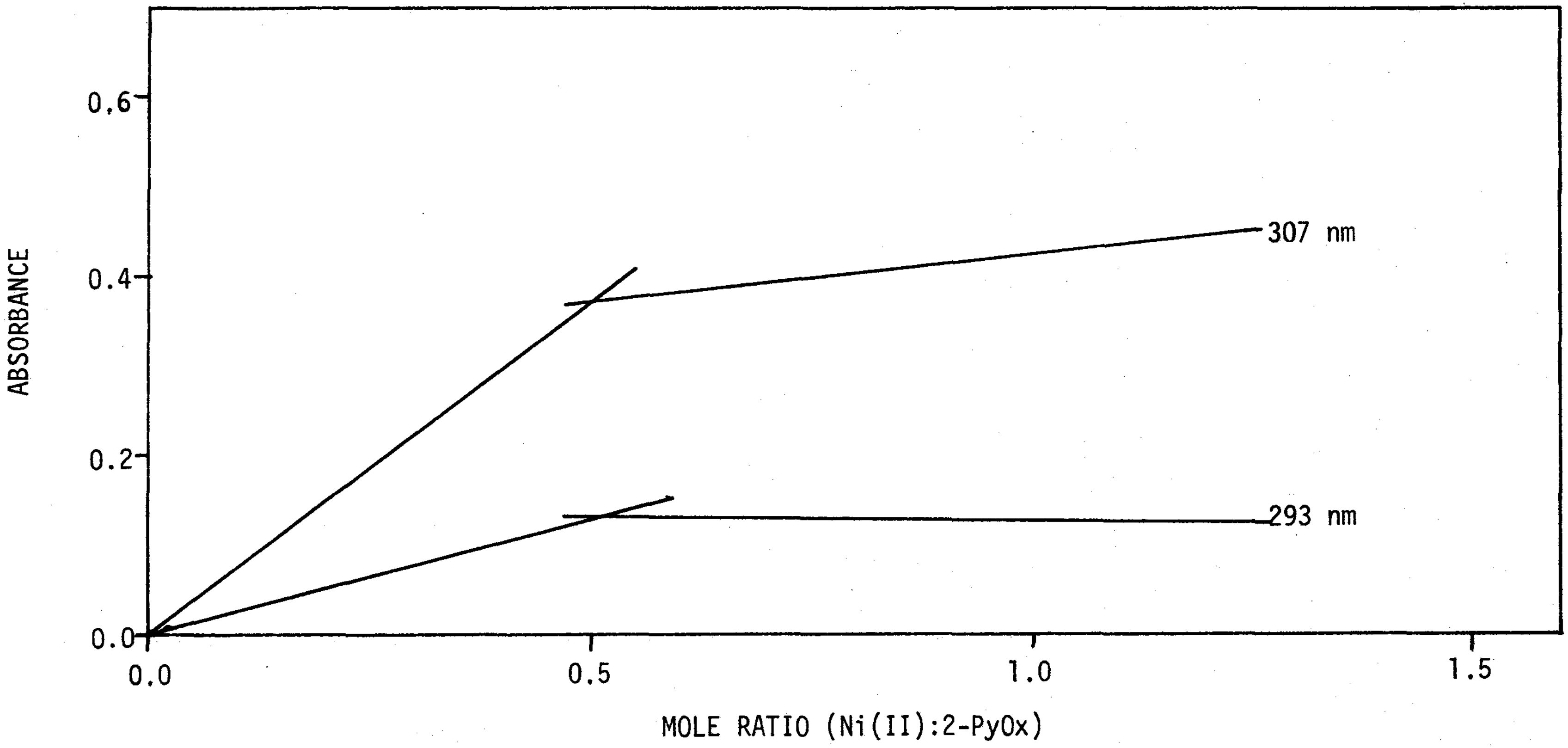


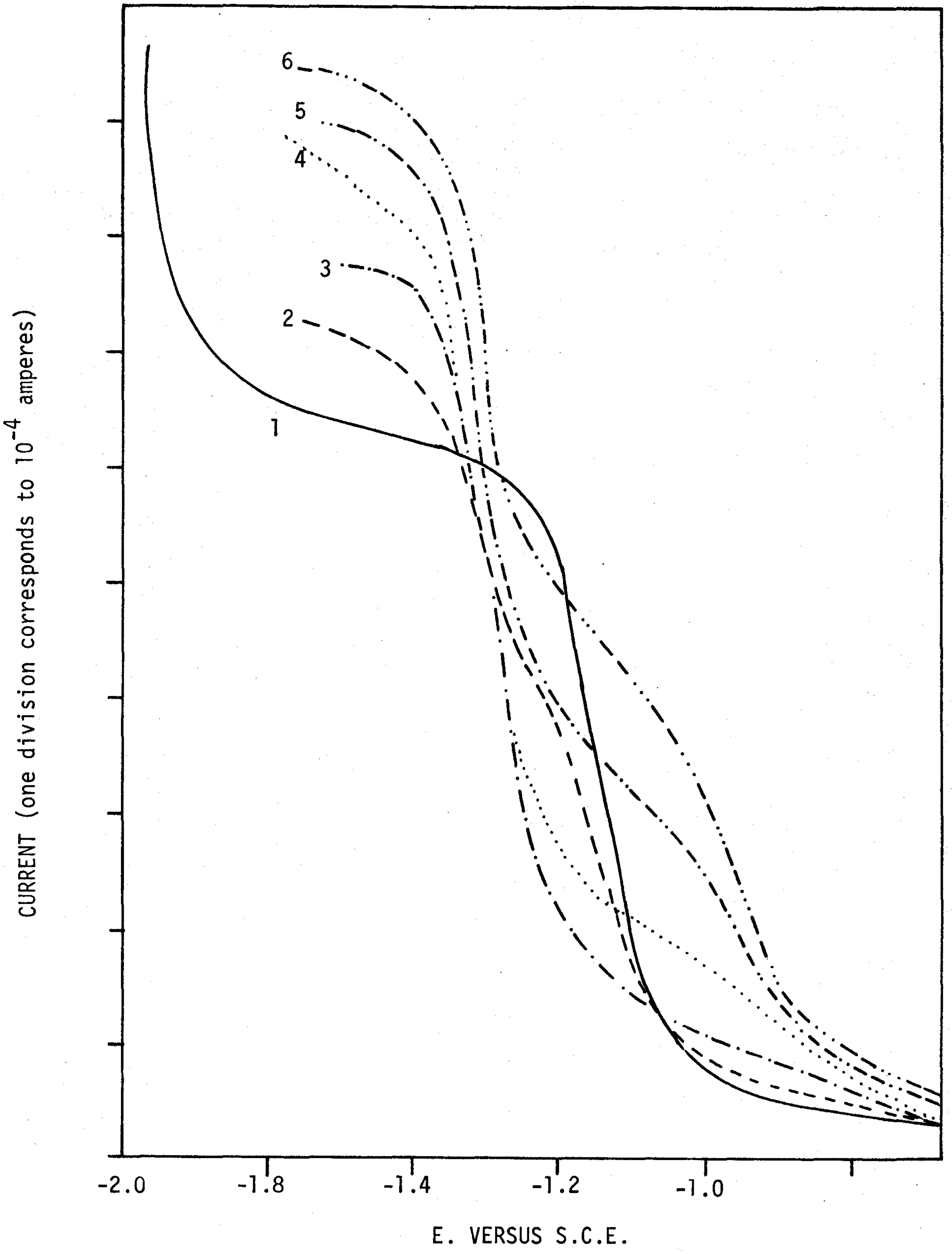
Figure 28. MOLE RATIO PLOTS FOR NICKEL(II) CHELATES OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE AT pH 5. [2-PyOx] = 2.0 x 10^{-5} M.

ratios. Since $\bar{n} = 0.5$ this species must be $NiLHL^+$.

Additional evidence for the formation of $NiLHL^+$ was obtained polarographically. Figure 29 shows polarograms recorded at pH 5 for a series of solutions which contained a constant concentration of 2-PyOx and varying concentrations of Ni(II). Data obtained from the polarograms were only semi-quantitative because of incomplete removal of oxygen from the solutions and because of the small difference in the half-wave potentials between Ni(II) ($E_{1/2} = -1.15$ v vs S.C.E.) and 2-PyOx ($E_{1/2} = -1.35$ v vs S.C.E.). However, some definite conclusions about the nature of the complex formed could be made. The polarogram of 2-PyOx at pH 5 is given by curve 1. With increasing Ni(II) concentration the portion of the polarogram corresponding to free 2-PyOx decreased, and at a 2-PyOx: Ni(II) mole ratio of 2.5:1 (curve 3) most of the free 2-PyOx had reacted with Ni(II). At a molar ratio of 1.7:1(curve 4) the current at $E = -1.2$ v increased. As the current at this voltage is the diffusion current of Ni(II), uncomplexed Ni(II) must be present in the solution. This diffusion current increased as the molar ratio was reduced further (i.e., as $[Ni^{2+}]$ was increased). The results show that free Ni(II) does exist at pH 5 when the molar ratio is in the range of 1:1 to 2:1. This conclusion is consistent with the formation of $NiLHL^+$ in preference to NiL^+.

Unlike the 2-PyOx/Cu(II) system, the preliminary investigation of the potentiometric data (in particular the formation curves) did not indicate whether $NiHL^{2+}$ and NiL^+ were formed in addition to $NiLHL^+$ and NiL_2; however, data from as many as eight different titrations were refined at one time with the program SCOGS and it was found that all five of the above complexes were required for an adequate description of this system.

Figure 29. POLAROGRAMS OF NICKEL(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE IN 50% v/v AQUEOUS DIOXANE AT pH 4.9. (1) 5.56 x 10^{-5} moles 2-PyOx in 118 ml (——), (2) as in 1 plus 1.15 x 10^{-5} moles Ni(II) in 120 ml (----), (3) as in 1 plus 2.30 x 10^{-5} moles Ni(II) in 122.5 ml (-·-·), (4) as in 1 plus 3.45 x 10^{-5} moles Ni(II) in 124.6 ml (····), (5) as in 1 plus 4.60 x 10^{-5} moles Ni(II) in 127.4 ml (---·), (6) as in 1 plus 5.75 x 10^{-5} moles Ni(II) in 130 ml (···--).


6
5
4
3
2
1
CURRENT (one division corresponds to 10^{-4} amperes)
-2.0
-1.8
-1.4
-1.2
-1.0
E. VERSUS S.C.E.

Reactions of 2-PyOx with Co(II), Zn(II) and Cd(II). The titration curves for these metal ions in the presence of 2-PyOx are given in Figures 30 - 32. A comparison of these titration curves with those for Ni(II) (Figure 25) shows that these metal ions react with 2-PyOx in a similar manner. For titrations of the 1:1 systems, the buffer zone in the pH region 5.5 - 7.0 was not as distinct as it was for the Ni(II) system. When larger concentrations of Co(II) and Zn(II) were used, this buffer zone was more prominent. The formation curves for these metal ions are also similar to those for the Ni(II) system and representative examples are given in Figure 33. The formation curve for Cd(II) indicates that the protonated complexes of Cd(II) are not as stable as those for Ni(II), Co(II) and Zn(II). An analysis of the titration data with SCOGS showed that MHL^{2+}, $M(HL)_2^{2+}$, ML^+, ML_2 and $MLHL^+$ were required for an adequate description of all three 2-PyOx/metal-ion systems.

Reactions of 2-PyOx with Al(III). Two titration curves for the 2-PyOx/Al(III) system are shown in Figure 34. These were performed at the molar ratios of 4:1 (curve 2) and 2:1 (curve 3). At pH ∿ 5 each curve was displaced from the $HClO_4$ curve by an amount corresponding to the liberation of two protons per metal ion. This displacement indicates the formation of AlL_2^+. A second buffer zone corresponding to the consumption of one mole of base per mole of metal ion occurred between 5.1 and 5.5 ml of base (pH 5.3 - 6.3). In this pH region the yellow color of the solution faded and a white flocculent precipitate*, characteristic of $Al(OH)_3 \cdot XH_2O$, appeared. The decrease in the pH of this buffer zone with a decrease in reagent

---

* The precipitate was isolated and found to contain essentially no carbon.

Figure 30. TITRATION CURVES OF COBALT(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. (1) 5.16×10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——), (2) as in 1 plus 3.65×10^{-4} moles 2-PyOx and 5.23×10^{-5} moles Co(II) (····), (3) as in 1 plus 1.06×10^{-4} moles 2-PyOx and 5.23×10^{-5} moles Co(II) (-·-·), (4) as in 1 plus 5.48×10^{-5} moles 2-PyOx and 5.23×10^{-5} moles Co(II) (----), (5) as in 1 plus 30 ml 50% (v/v) aqueous dioxane and 2.09×10^{-4} moles 2-PyOx and 2.09×10^{-4} moles 2-PyOx and 2.09×10^{-4} moles Co(II) (---·).

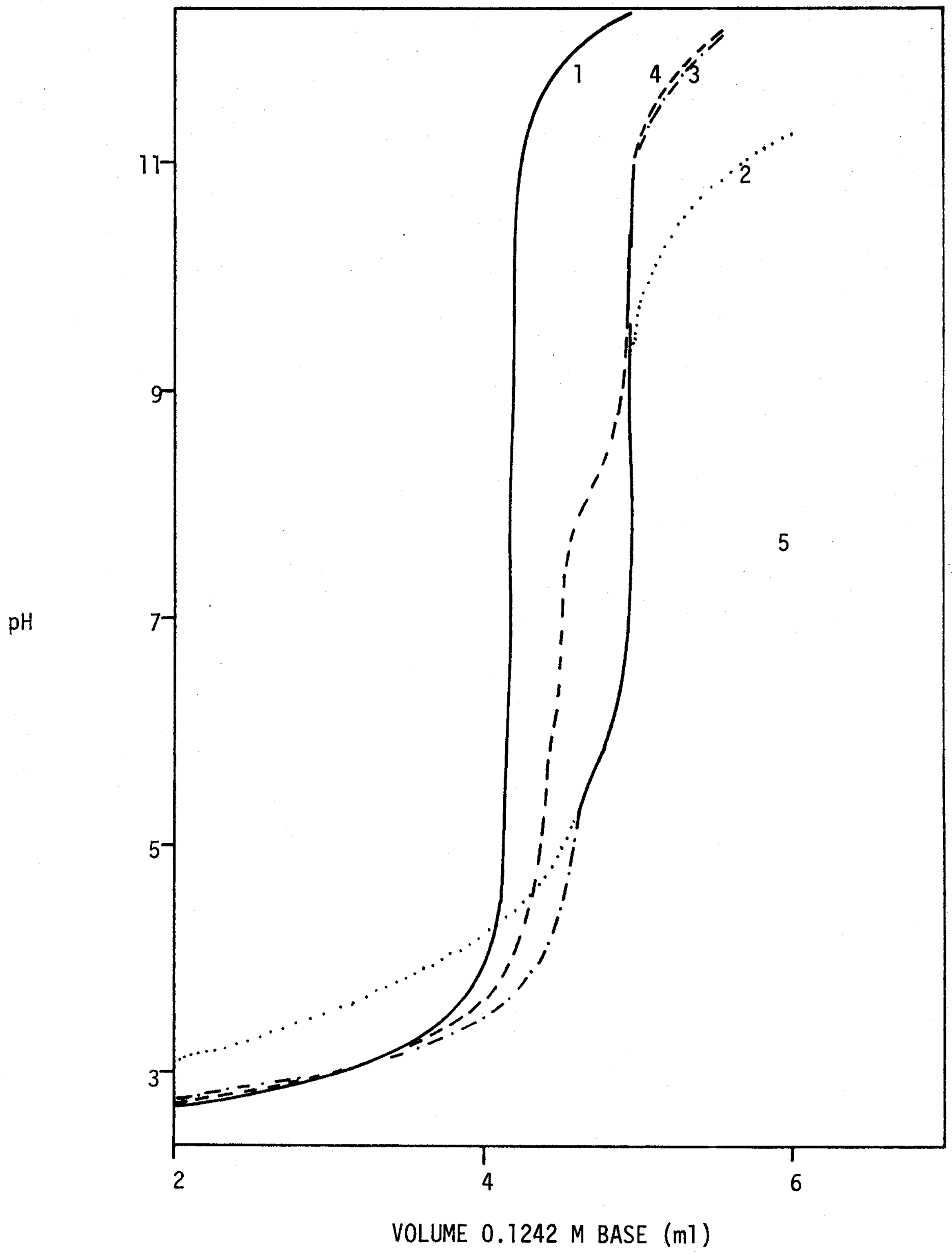

1
4
3
2
5
pH
11
9
7
5
3
2
4
6
VOLUME 0.1242 M BASE (ml)

Figure 31. TITRATION CURVES OF ZINC(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. (1) 5.16×10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——), (2) as in 1 plus 2.49×10^{-4} moles 2-PyOx and 5.56×10^{-5} moles Zn(II) (····), (3) as in 1 plus 1.13×10^{-4} moles 2-PyOx and 5.56×10^{-5} moles Zn(II) (-·-·), (4) as in 1 plus 5.63×10^{-5} moles 2-PyOx and 5.56×10^{-5} moles Zn(II) (----).

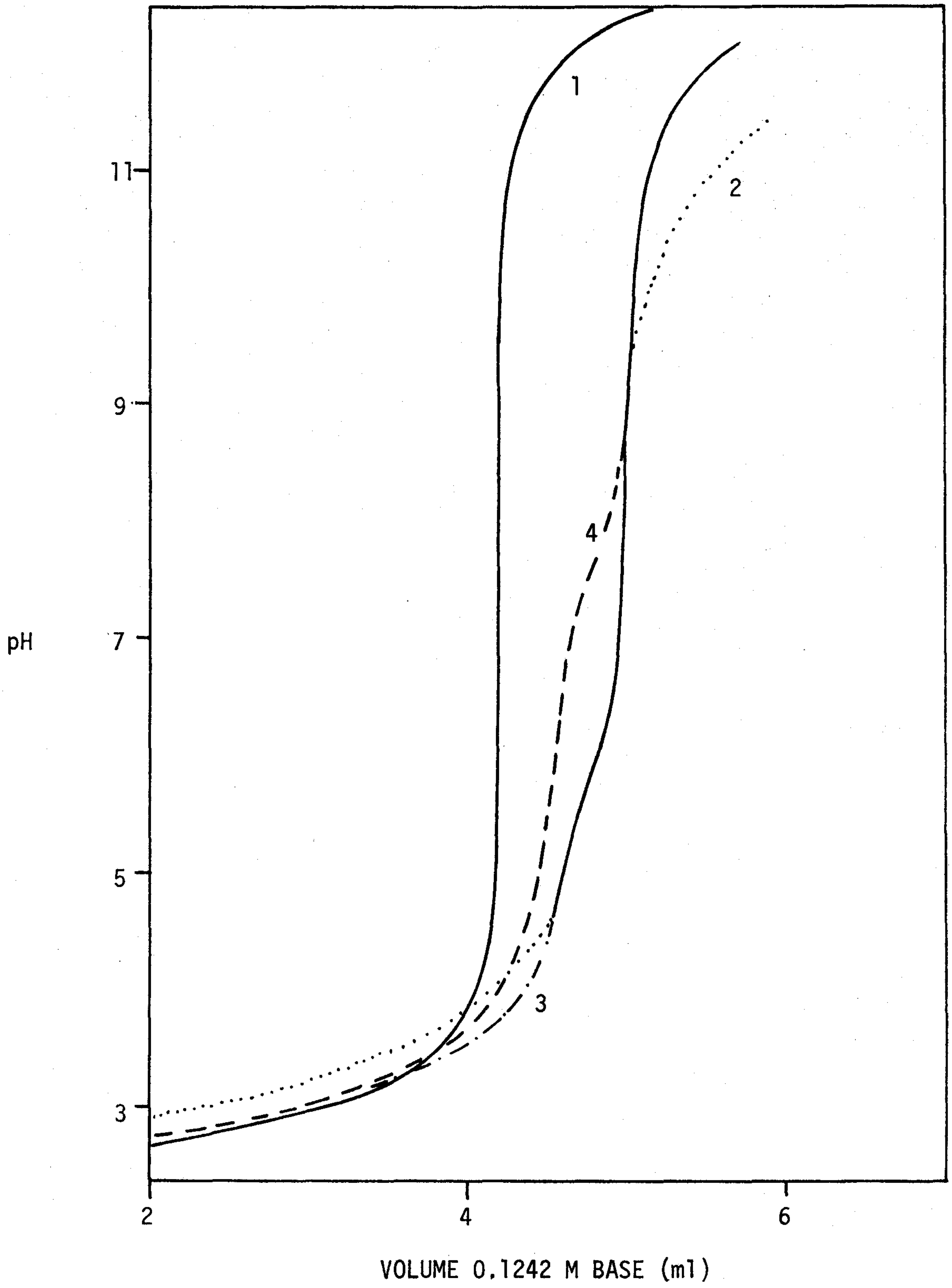
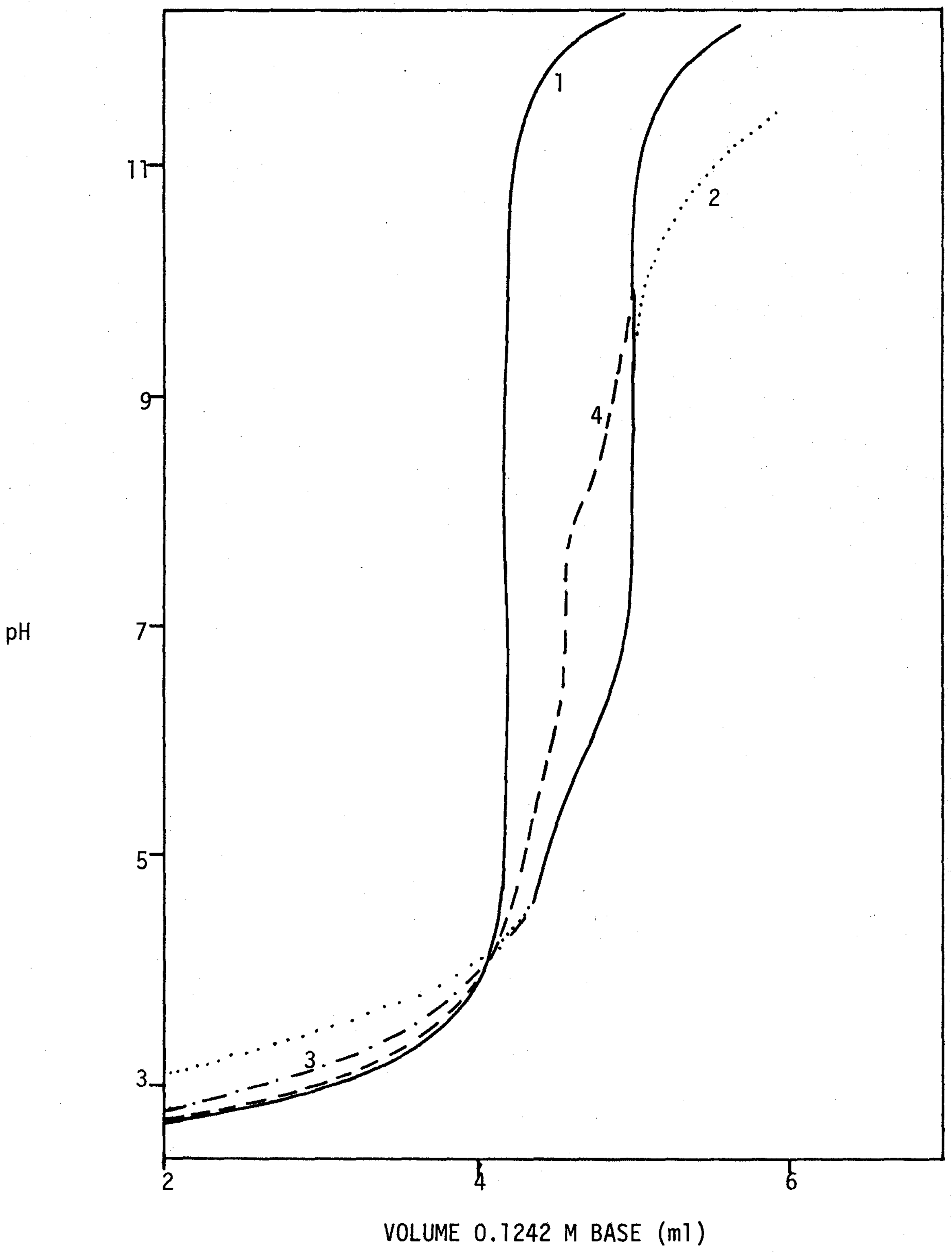

pH
3
5
7
9
11
2
4
6
VOLUME 0.1242 M BASE (ml)
1
2
3
4

Figure 32. TITRATION CURVES OF CADMIUM(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. (1) 5.16×10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——), (2) as in 1 plus 2.63×10^{-4} moles 2-PyOx and 5.46×10^{-5} moles Cd(II) (····), (3) as in 1 plus 1.10×10^{-4} moles 2-PyOx and 5.46×10^{-5} moles Cd(II) (-·-·), (4) as in 1 plus 5.54×10^{-5} moles 2-PyOx and 5.46×10^{-5} moles Cd(II) (----).


pH
11
9
7
5
3
1
2
4
3
2
4
6
VOLUME 0.1242 M BASE (ml)

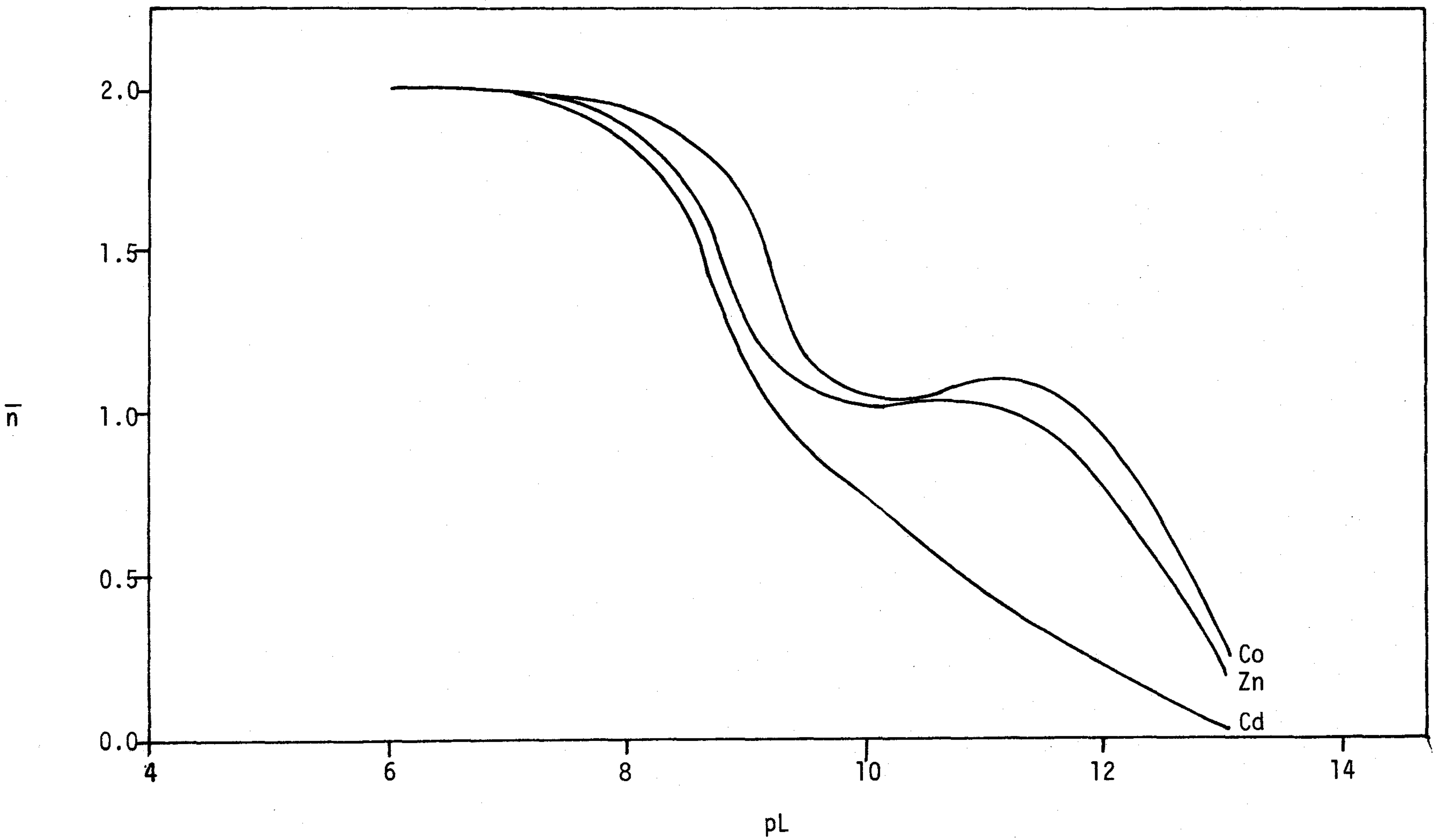
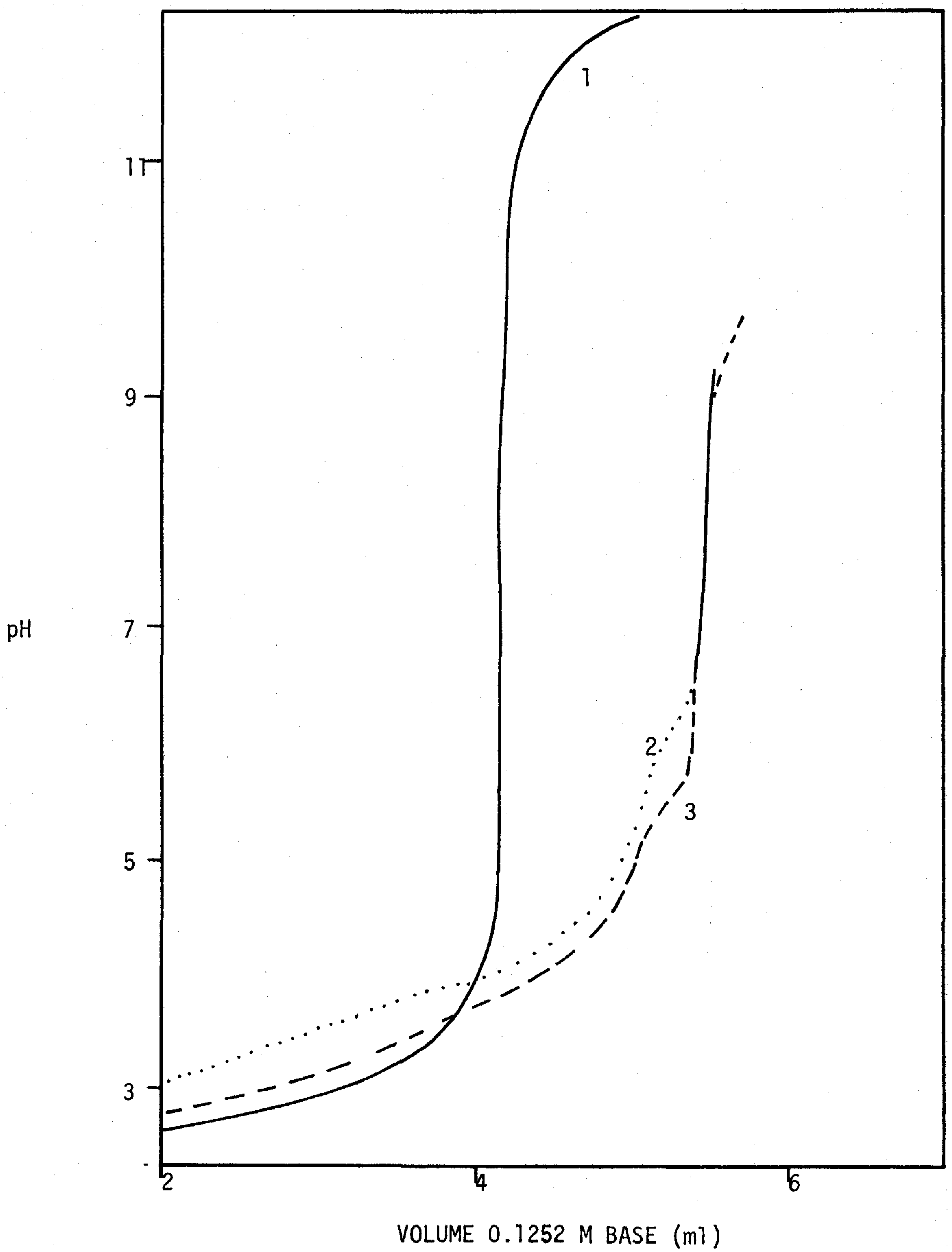


Figure 33. FORMATION CURVES FOR COBALT(II), ZINC(II), AND CADMIUM(II) CHELATES OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE. The molar ratio of 2-PyOx:metal ion = 2:1.

Figure 34. TITRATION CURVES OF ALUMINUM(III) AND 2-(2'PYRIDYL)-8-HYDROXYQUINOLINE. (1) 5.21×10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——), (2) as in (1) plus 2.24×10^{-4} moles 2-PyOx and 5.34×10^{-5} moles Al(III) (····), (3) as in 1 plus 1.12×10^{-4} moles 2-PyOx and 5.34×10^{-5} moles Al(III) (----).


pH
3
5
7
9
11
2
4
6
VOLUME 0.1252 M BASE (ml)
1
2
3

concentration provided further evidence of hydrolysis. A distortion in the initial portion of the formation curve showed that partial hydrolysis occurred prior to chelate formation. As a result reliable formation constants could not be calculated. An approximate value of log K(AlL) ≈ 10 was obtained from the formation curve. This value is considerably lower than that for the corresponding complex of 2-PhOx (log K(AlL) = 19.8). The facile hydrolysis of the $\mathrm{Al(2\text{-}PyOx)_2^+}$ complex compared to the $\mathrm{Al(2\text{-}PhOx)_2^-}$ complex is further evidence of reduced stability. The lower stability is undoubtedly related to reduced charge neutralization on chelate formation, lower donor-atom basicity, and the "softer" nature of the nitrogen donor atom versus the oxygen donor atom(15).

Solid Chelates of 2-PyOx. Attempts were made to prepare solid 2-PyOx chelates of each of the metal ions. A solid Al(III) chelate could not be isolated. For the solid chelates isolated, the results of C, H and N elemental analysis are given in TABLE VII. The Cu(II) and Cd(II) chelates could not be obtained in a pure state.

Formation Constants for 2-PyOx Chelates. The formation constants for the 2-PyOx chelates are listed in TABLE VIII. A corrected and modified version of the program SCOGS (APPENDIX II) was used to calculate these constants.

The standard deviation in titre (i.e., volume) of base (σ_T), TABLE VIII, is a measure of the agreement between the titre of base which is calculated using the refined formation constants, and the experimental titre. It gives the "goodness of fit" for the calculated constants and its value is in the range of 20 - 40 x 10^{-3} ml. The value for Cu(II) is

TABLE VII

ELEMENTAL ANALYSIS FOR METAL CHELATES OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE

	%C	%H	%N
Calculated for $Ni(C_{14}H_9ON_2)_2$:	67.1	3.6	11.2
Found:	65.6	3.8	11.3
Calculated for $Co(C_{14}H_9ON_2)2$:	67.1	3.6	11.2
Found:	66.1	3.8	11.3
Calculated for $Zn(C_{14}H_9ON_2)_2$:	66.2	11.0	3.6
Found:	65.6	11.1	3.7
Calculated for $Ni(C_{14}H_9ON_2)(C_{14}H_{10}ON_2)ClO_4$:	55.0	9.3	3.2
Found:	54.7	9.3	3.4

TABLE VIII

FORMATION CONSTANTS OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE METAL CHELATES

(50% (v/v) aqueous dioxane, 25.0°C, ionic strength = 0.1)

Metal Ion	$\beta(ML)$*	σ**	$\beta(ML_2)$	σ	$\beta(MHL)$	σ	$\beta(MH_2L_2)$	σ	$\beta(MHL_2)$	σ	$\sigma_T \times 10^{-3}$+	No. Titns.++
Mn(II)	8.442	0.007	16.496	0.004	—	—	—	—	—	—	3.1	1
Ni(II)	11.8	0.2	24.03	0.1	16.62	0.09	32.8	0.1	30.40	0.09	35	3
Co(II)	11.07	0.09	22.42	0.03	15.50	0.05	29.8	0.2	28.11	0.03	14	3
Cu(II)	14.65	0.08	22.75	0.09	—	—	—	—	30.06	0.07	50	5
Zn(II)	10.89	0.2	21.61	0.03	14.64	0.2	29.85	0.07	27.46	0.02	26	3
Cd(II)	9.1	0.4	19.60	0.03	14.68	0.04	29.11	0.03	25.64	0.02	18	3
Ni(II)‡	11.2	0.1	22.43	0.06	15.48	0.08	30.60	0.06	28.36	0.06	24	3

---

*All formation constants are in log units. $\beta(MH_iL_j) = [MH_iL_j]/([M][H]^i[L]^j)$

**σ is the estimated standard deviation for the preceding constant.

+σ_T is the standard deviation in titre.

++The number of titrations refined together.

‡Data for 60.0°C.

outside this range since the titration data were obtained over a period of seven months. When both the number of species formed and the number of titrations are reduced the value of σ_T becomes quite small, as is shown by the value of σ_T for Mn(II) (the Mn(II) complexes are discussed in a later section).

An advantage of SCOGS and related programs is that the least squares refinement employed by such programs gives an estimate of the standard deviation, σ , for each formation constant. The magnitude of the value σ is determined by the "closeness of fit" between the calculated constants and the experimental data. The reliability of σ for a particular constant depends on the concentration of the corresponding complex relative to the total metal-ion concentration, and on the number of titrations refined. A realistic estimate of σ for a major species can be obtained if three titrations are refined together. For a minor species, however, this is not necessarily the case. The concentration of a minor species may be significant for only a small portion of the data. Consequently, the calculated value of the formation constant for such a species can be significantly affected by small experimental errors. This follows since the introduction of a large error in the formation constant could give a smaller value of σ_T which in turn will give a smaller value of σ for the constant. Thus, before any significance is attached to the value of a formation constant it is necessary to know whether the complex is a major or minor component. This can be determined from a plot of α_A (fraction of the total metal-ion concentration present as the complex A) versus pH. The values of α_A can be calculated from the refined formation constants.

Figures 35, 36 and 37 show such α_A plots for three 2-PyOx/Ni(II)

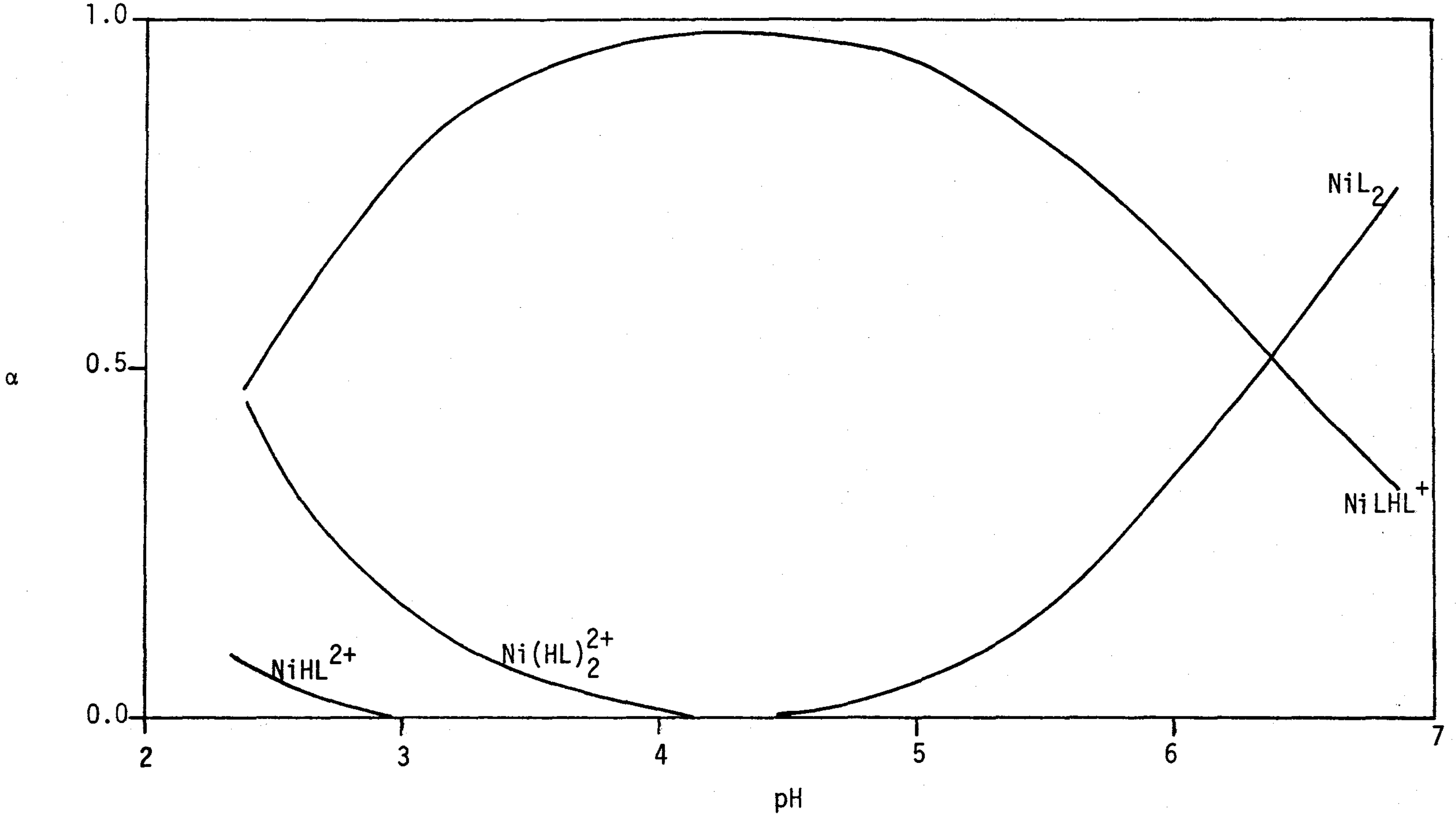


Figure 35. α CURVES FOR THE TITRATION OF NICKEL(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE IN A 4:1 MOLAR RATIO.

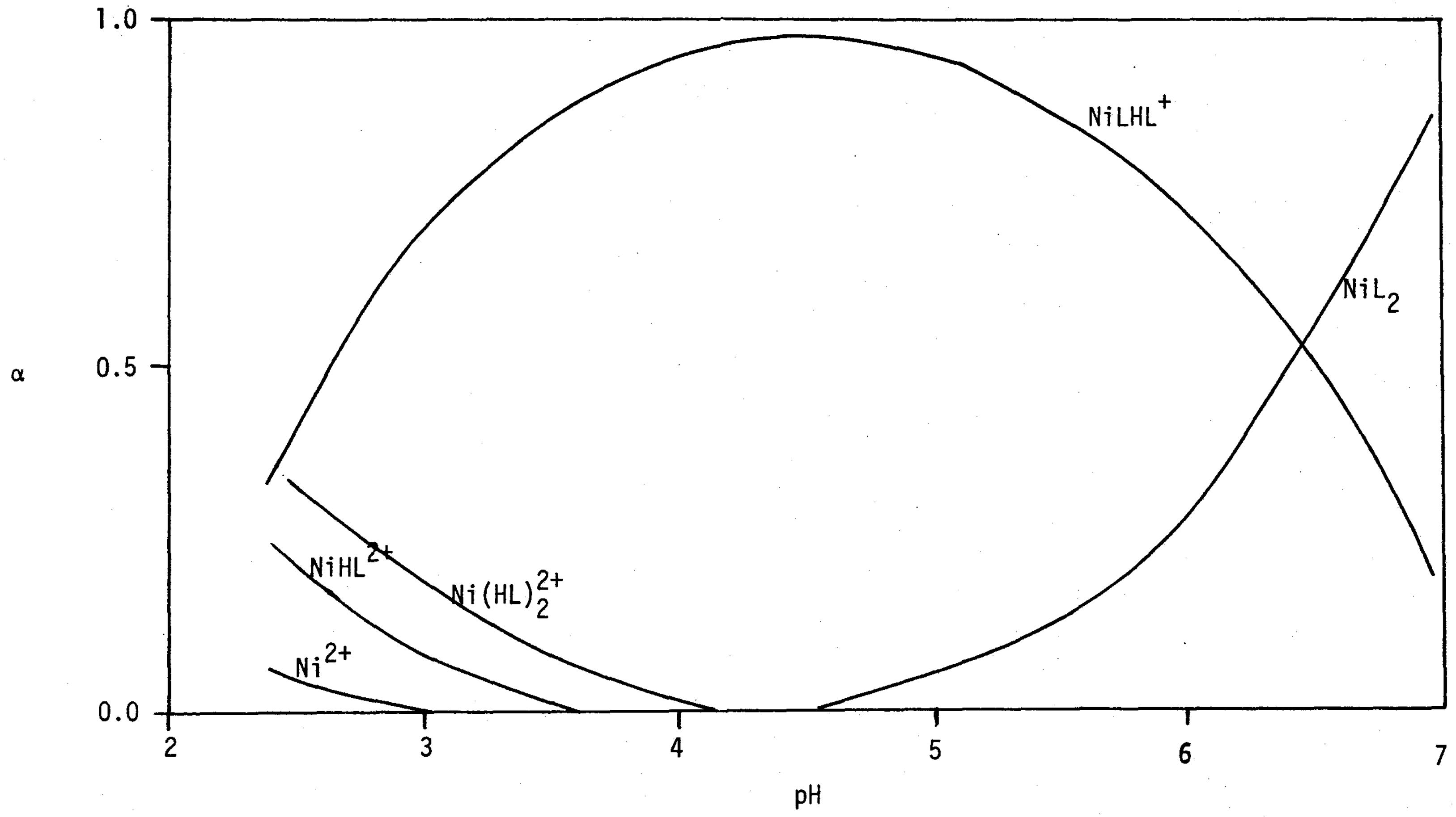


Figure 36. α CURVES FOR THE TITRATION OF NICKEL(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE IN A 2:1 MOLAR RATIO.

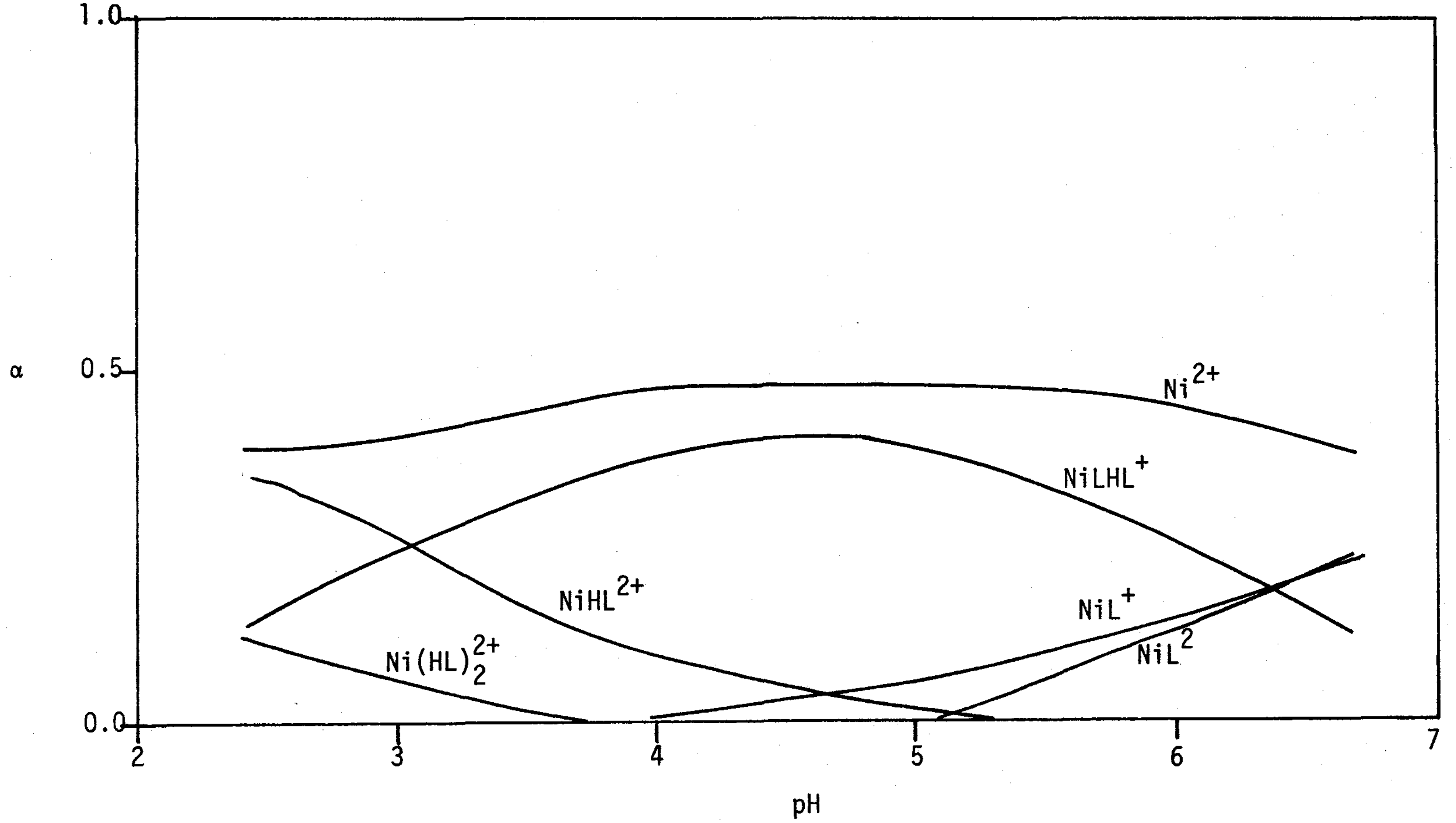


Figure 37. α CURVES FOR THE TITRATION OF NICKEL(II) AND 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE IN A 1:1 MOLAR RATIO.

titrations in molar ratios of 4:1, 2:1 and 1:1 respectively. These curves show that at the lower pH values the species $NiHL^{2+}$ and $Ni(HL)_2^{2+}$ are predominant; in the intermediate range (pH 3 - 6) $NiLHL^+$ is predominant, and at pH 7, NiL_2 is predominant. The α_A plots for the Co(II), Zn(II) and Cd(II) systems were similar and are not shown.

Where a large number of complexes are required for an adequate description of the system it is imperative that the concentrations of all the species present in the pH range over which data have been obtained be calculated with the refined constants and then compared with the results obtained from the titration and formation curves. This comparison is best done with the use of α_A plots. This procedure was used for each of the 2-PyOx/metal-ion systems. In each case the concentrations calculated with the formation constants agreed with the results of the formation and titration curves. An interesting aspect of these α_A plots is that they showed that a special stability is associated with $MLHL^+$ for Ni(II), Co(II), Zn(II) and Cd(II) but not for Cu(II). The value of $\beta(MLHL)$ is greatest for the Cu(II) chelate, but Cu(II) is the only one of the above metal ions which forms ML^+ prior to the formation of $MLHL^+$. This behavior of Cu(II) will be discussed later in a separate section.

The reactions shown below summarize the results found for the reactions of 2-PyOx with Co(II), Ni(II), Zn(II) and Cd(II). Cu(II) forms only CuL^+, $CuLHL^+$ and CuL_2 in significant amounts. The approximate pH ranges given only apply to the results of this study. The ranges will change with changes in experimental conditions such as concentrations of reacting species.

$$M^{2+} + H_2L^+ \rightleftarrows H^+ + MHL^{2+} \qquad \text{pH 2.5 - 3.5}$$

$MHL^{2+} + H_2L^+ \rightleftarrows H^+ + M(HL)_2^{2+}$ pH 2.5 - 3.5

$M^{2+} + H_2L^+ \rightleftarrows 2H^+ + ML^+$ (small amounts) pH 2.5 - 3.5

$ML^+ + H_2L^+ \rightleftarrows H^+ + MLHL^+$ pH 2.5 - 5.0

$MLHL^+ \rightleftarrows H^+ + ML_2$ pH 5.0 - 7.0

Metal Chelates of 2-Hydrazino-8-hydroxyquinoline

The reactions of 2-HyOx with Co(II), Ni(II), Cu(II), Zn(II) and Cd(II), based on potentiometric data, are discussed in this section. The metal-chelate formation constants have been calculated and are listed in TABLE IX. No reaction was observed with Al(III).

Reaction of 2-HyOx with Co(II), Ni(II), Cu(II), Zn(II) and Cd(II).

The titration curves for 2-HyOx alone and for the 2-HyOx/metal-ion systems are shown in Figures 38 - 43. These titrations were performed at molar ratios of 1:1 (curve 4), 2:1 (curve 3) and 5:1 (curve 2).

A comparison of the titration curves for the 2-HyOx/metal-ion systems in 2:1 and 5:1 molar ratios with those for free 2-HyOx shows that three protons per metal ion were liberated below pH 5.5. Since 2-HyOx is present in solution mainly as H_2L^+ at pH 5.5, the above results suggest the formation of $MLHL^+$. This behavior is similar to that found previously for 2-PyOx systems.

The acid dissociation of the unreacted H_2L^+ obscures the buffer zone for the formation of ML_2; however, at pH 9, where 2-HyOx is present mainly as HL, the titration curves for 2:1 and 5:1 molar ratio systems were shifted from the $HClO_4$ curve by an amount corresponding to the release of two protons per metal ion. This fact is consistent with the formation

Figure 38. TITRATION CURVES OF 2-HYDRAZINO-8-HYDROXYQUINOLINE. (1) 5.21×10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——), (2) as 1 plus 5.72×10^{-5} moles 2-HyOx (----), (3) as 1 plus 1.13×10^{-4} moles 2-HyOx (-·-·), (4) as 1 plus 2.84×10^{-4} moles 2-HyOx (····).

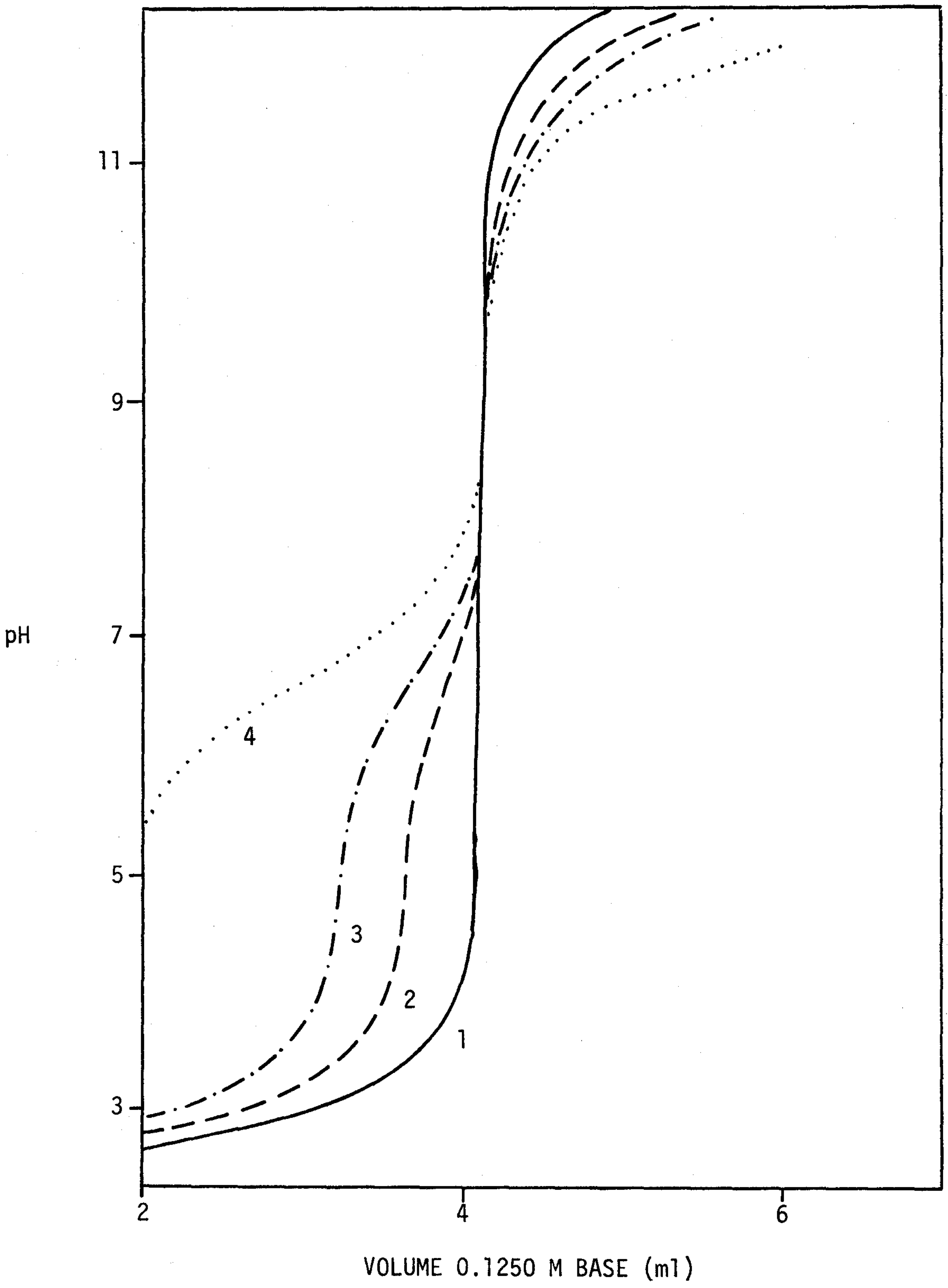

pH
11
9
7
5
3
2
4
6
VOLUME 0.1250 M BASE (ml)
1
2
3
4

Figure 39. TITRATION CURVES OF COBALT (II) AND 2-HYDRAZINO-8-HYDROXYQUINOLINE. (1) 5.21×10^{-2} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——),
(2) as 1 plus 2.62×10^{-4} moles 2-HyOx and 5.23×10^{-5} moles Co(II) (····),
(3) as 1 plus 1.07×10^{-4} moles 2-HyOx and 5.23×10^{-5} moles Co(II) (-·-·),
(4) as 1 plus 5.03×10^{-5} moles 2-HyOx and 5.23×10^{-5} moles Co(II) (----).

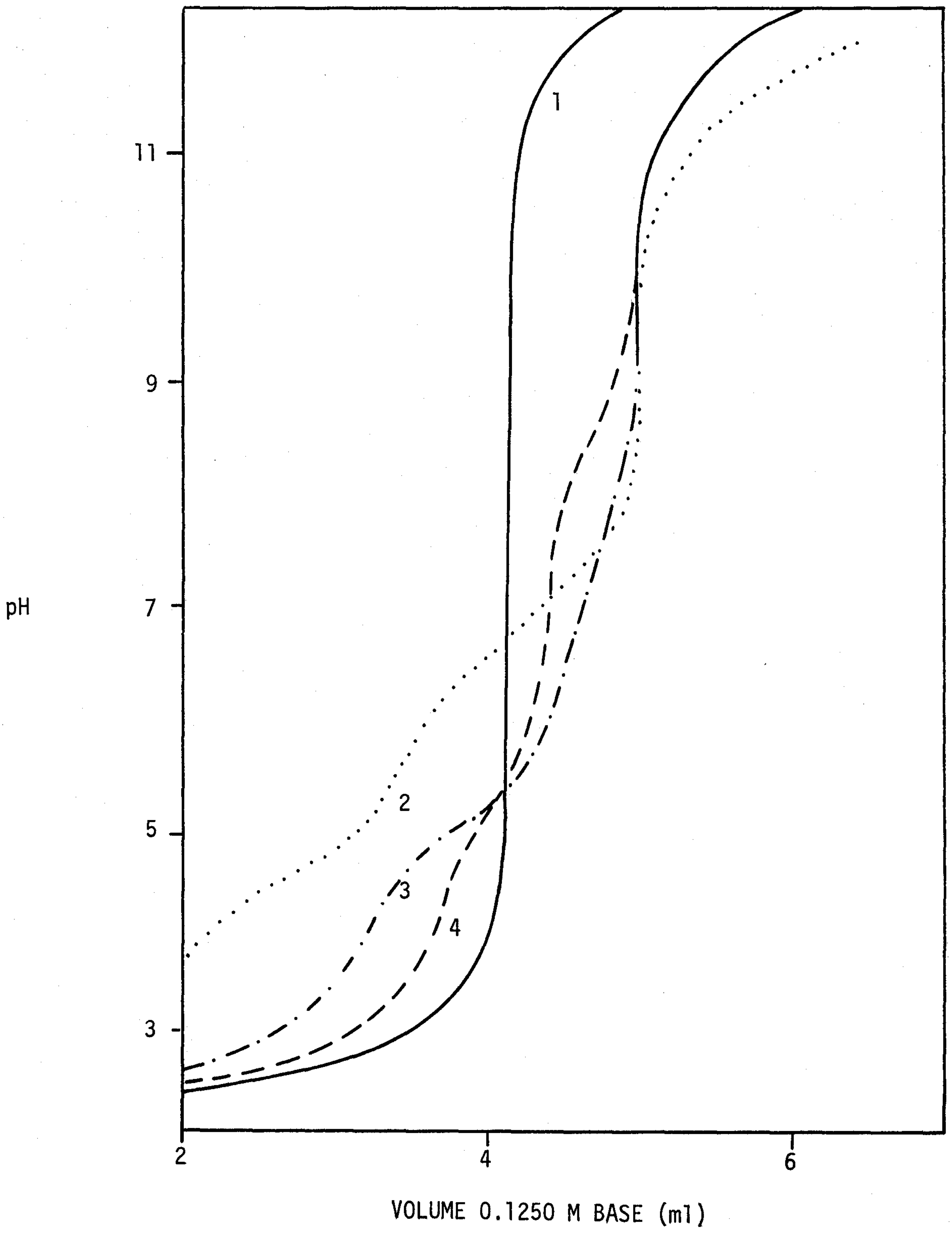

pH
11
9
7
5
3
2
4
6
VOLUME 0.1250 M BASE (ml)
1
2
3
4

Figure 40. TITRATION CURVES OF NICKEL (II) AND 2-HYDRAZINO-8-HYDROXYQUINOLINE.
(1) 5.21×10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——),
(2) as 1 plus 2.84×10^{-4} moles 2-HyOx and 5.76×10^{-5} moles Ni(II) (····),
(3) as 1 plus 1.16×10^{-4} moles 2-HyOx and 5.76×10^{-5} moles Ni(II) (-·-·),
(4) as 1 plus 5.35×10^{-5} moles 2-HyOx and 5.76×10^{-5} moles Ni(II) (----).

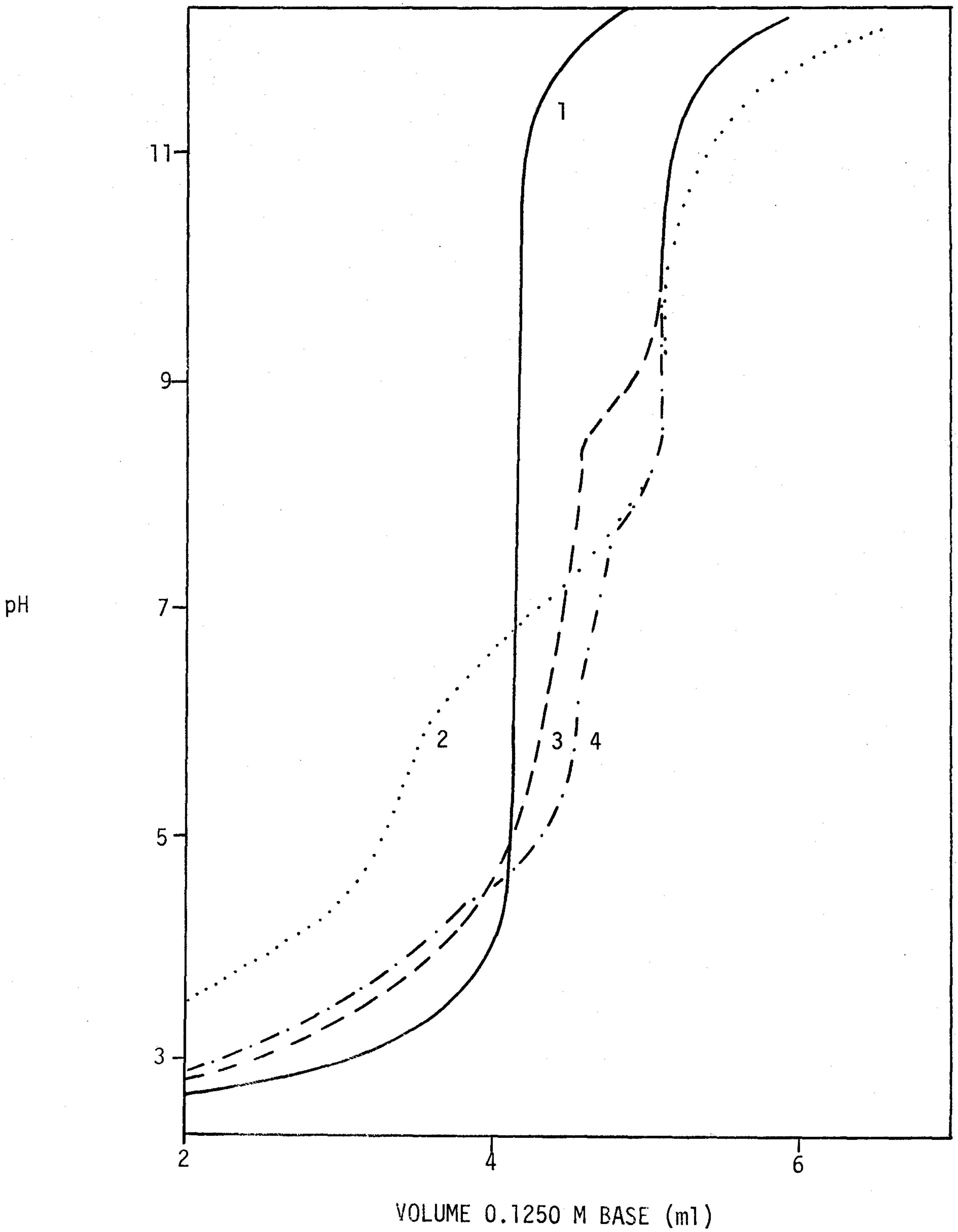

pH
11
9
7
5
3
2
4
6
1
2
3
4
VOLUME 0.1250 M BASE (ml)

Figure 41. TITRATION CURVES OF COPPER(II) AND 2-HYDRAZINO-8-HYDROXYQUINOLINE.
(1) 5.21 x 10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——),
(2) as 1 plus 2.88 x 10^{-4} moles 2-HyOx plus 5.40 x 10^{-5} moles Cu(II) (····),
(3) as 1 plus 1.15 x 10^{-4} moles 2-HyOx and 5.40 x 10^{-5} moles Cu(II) (-···),
(4) as 1 plus 5.83 x 10^{-5} moles 2-HyOx and 5.40 x 10^{-5} moles Cu(II) (----).

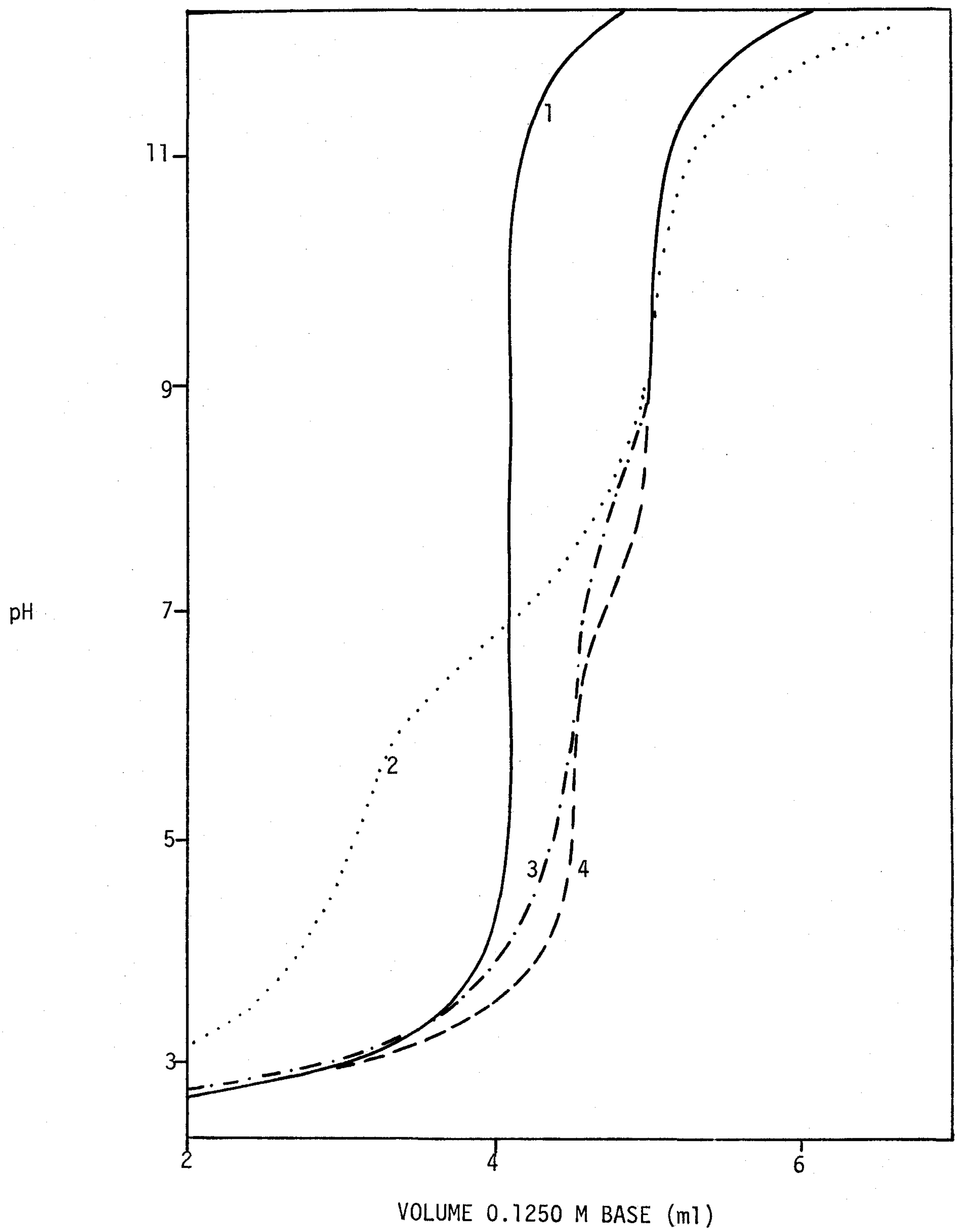

pH
11
9
7
5
3
2
4
6
VOLUME 0.1250 M BASE (ml)
1
2
3
4

Figure 42. TITRATION CURVES OF ZINC(II) AND 2-HYDRAZINO-8-HYDROXYQUINOLINE.
(1) 5.21 x 10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——),
(2) as 1 plus 2.72 x 10^{-4} moles 2-HyOx and 5.56 x 10^{-5} moles Zn(II) (····),
(3) as 1 plus 1.13 x 10^{-4} moles 2-HyOx and 5.56 x 10^{-5} moles Zn(II) (-·-·),
(4) as 1 plus 6.06 x 10^{-5} moles 2-HyOx and 5.56 x 10^{-5} moles Zn(II) (----).

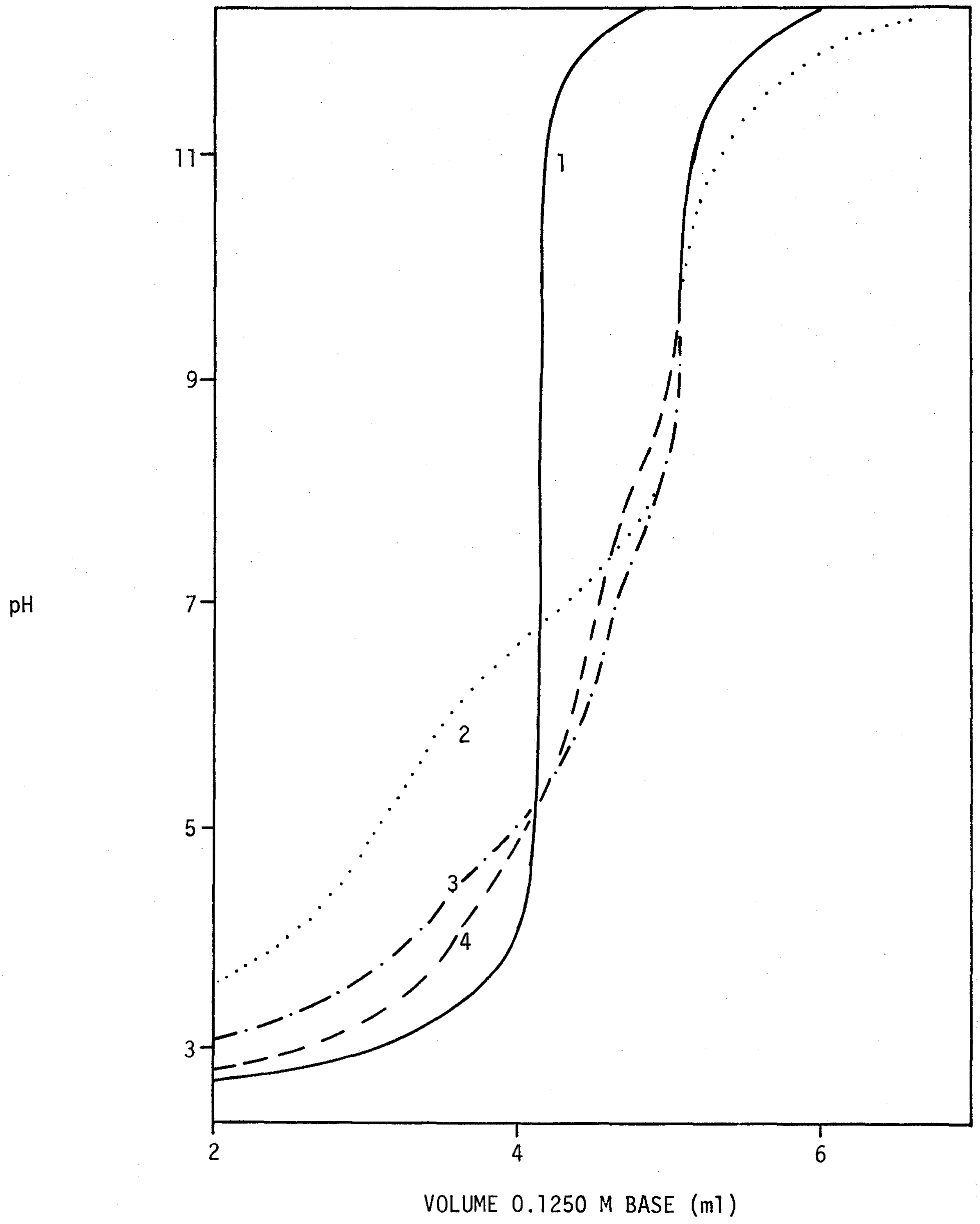

pH
11
9
7
5
3
1
2
3
4
2
4
6
VOLUME 0.1250 M BASE (ml)

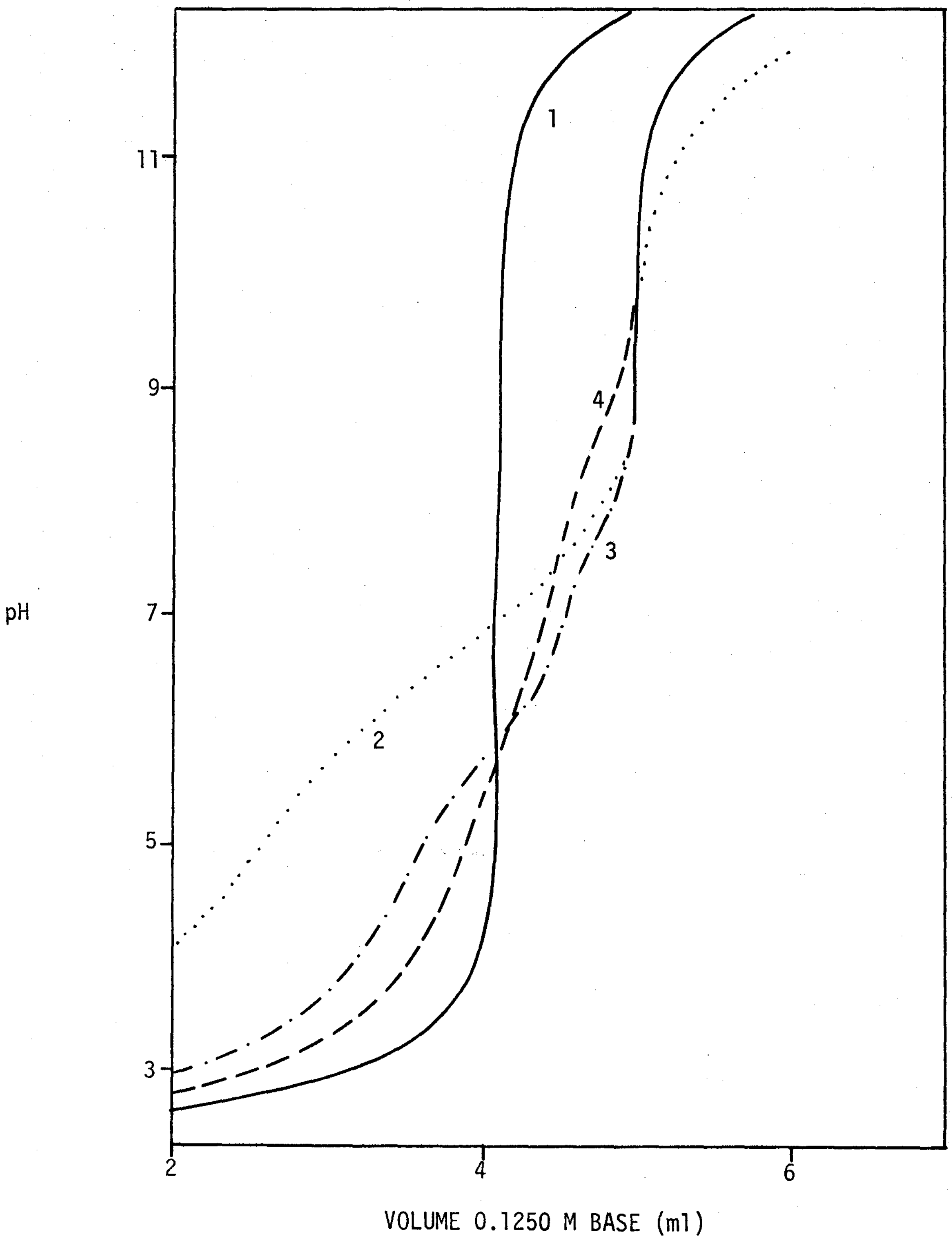
Figure 43. TITRATION CURVES OF CADMIUM(II) AND 2-HYDRAZINO-8-HYDROXYQUINOLINE.

(1) 5.21 x 10^{-4} moles $HClO_4$ in 110 ml 50% (v/v) aqueous dioxane (——),

(2) as 1 plus 2.84 x 10^{-4} moles 2-HyOx and 5.46 x 10^{-5} moles Cd(II) (....),

(3) as 1 plus 1.09 x 10^{-4} moles 2-HyOx and 5.46 x 10^{-5} moles Cd(II) (-·-·),

(4) as 1 plus 6.08 x 10^{-5} moles 2-HyOx and 5.46 x 10^{-5} moles Cd(II) (----).


pH
11
9
7
5
3
2
4
6
VOLUME 0.1250 M BASE (ml)
1
2
3
4

of ML_2 at higher pH values. The titration curves for the 2-HyOx/Cd(II) system did not show any distinctive breaks prior to pH 9, which indicates that the Cd(II) complexes are not as stable as the other metal chelates.

The titrations for the Ni(II) system exemplifies the reactions found for 1:1 molar ratio titrations (Figure 40). In the Ni(II) system 1.5 protons per metal ion were liberated below pH 6 suggesting that the following reaction occurred:

$$1/2\ Ni^{2+} + H_2L^+ \rightleftarrows 1/2\ NiLHL^+ + 3/2\ H^+$$

According to this equation, one-half of the initial amount of Ni(II) remained uncomplexed. In this respect 2-HyOx is similar to 2-PyOx in forming $NiLHL^+$ without prior formation of NiL^+. The next buffer zone in this titration corresponded to the liberation of 0.5 moles of protons per mole of Ni(II), which is the amount required for the dissociation of the proton in $NiLHL^+$.

The formation curves* for the 2-HyOx/metal-ion systems are similar to those for the corresponding 2-PyOx systems and are not shown here. The above interpretation of the titration curves is consistent with the formation curves in that the latter also show that protonated complexes, in particular $MLHL^+$, are required for an adequate description of the 2-HyOx systems.

Analysis of the potentiometric data with SCOGS showed that the complexes MHL^{2+}, $M(HL)_2^{2+}$, $MLHL^+$, ML^+ and ML_2 were formed in all of the 2-HyOx/metal-ion systems.

---

*The values of $\bar{n}$ and pL were calculated using equations for oxine/metal-ion systems (APPENDIX II).

Formation Constants for 2-HyOx Chelates. The formation constants for the 2-HyOx chelates are listed in TABLE IX. The corrected and modified version of SCOGS was used to calculate the constants. These constants were used to construct α_A plots for each system. These α_A plots were consistent with the interpretation of the titration and formation curves. They are similar to those for the 2-PyOx/Ni(II) system (Figures 35 - 37) and are not shown here.

The reactions shown below summarize the results found for the 2-HyOx/Co(II), Ni(II), Cu(II), Zn(II) and Cd(II) systems.

$M^{2+} + H_2L^+ \rightleftarrows H^+ + MHL^{2+}$ pH 4 - 6

$MHL^{2+} + H_2L^+ \rightleftarrows H^+ + M(HL)_2^{2+}$ pH 4 - 6

$M^{2+} + H_2L^+ \rightleftarrows 2H^+ + ML^+$ (small amount) pH 4 - 5

$ML^+ + H_2L^+ \rightleftarrows H^+ + MLHL^+$ pH 4 - 7

$MLHL^+ \rightleftarrows H^+ + ML_2$ pH 6 - 8

The Non-Protonated Complexes, ML^+ and ML_2, of 2-PyOx and 2-HyOx

As shown in the previous section, 2-PyOx and 2-HyOx form both protonated and non-protonated complexes. The present discussion is limited to the non-protonated complexes. The protonated complexes are discussed in the next section.

The species ML^+ forms only to a minor extent. Its concentration is not significant at any pH value when the ligand:metal-ion molar ratio is $\geq$ 2:1 (e.g., see α_A plots, Figures 35 - 37). Its concentration is kept small by the formation of $MLHL^+$. The only reliable K(ML) value obtained

TABLE IX

FORMATION CONSTANTS OF 2-HYDRAZINO-8-HYDROXYQUINOLINE METAL CHELATES
(50% (v/v) aqueous dioxane, 25.0°C, ionic strength = 0.1)

Metal Ion	$\beta(ML)$[*]	σ[**]	$\beta(ML_2)$	σ	$\beta(MHL)$	σ	$\beta(MH_2L_2)$	σ	$\beta(MHL_2)$	σ	σ_T x 10^{-3}[+]	No. Titr.[++]
Co(II)	10.6	0.2	21.19	0.05	16.2	0.2	32.8	0.1	28.59	0.02	33	3
Ni(II)	11.3	0.2	22.72	0.06	16.9	0.2	34.79	0.04	30.48	0.03	36	3
Cu(II)	16.04	0.03	25.87	0.05	18.90	0.08	38.20	0.03	33.95	0.04	19	3
Zn(II)	10.	2.	21.84	0.09	16.8	0.8	34.59	0.05	29.35	0.07	50	3
Cd(II)	8.	2.	19.29	0.05	16.67	0.08	33.16	0.05	27.16	0.05	28	3
Mn(II)	8.81	0.01	16.46	0.01	-	-	-	-	-	-	8	1

---

[*] All constants are given in log units, and are defined as previously stated (TABLE VIII).

[**] σ is the estimated standard deviation for the preceding constant.

[+] σ_T is the standard deviation in titre.

[++] The number of titrations refined together.

was for CuL^+. Comparison of this value with the corresponding K(CuL) value for oxine shows that this species is terdentate (see TABLES VI, VIII and IX). The ML^+ species of the other metal ions are also likely terdentate.

The complex ML_2 arises from deprotonation of $MLHL^+$. For 2-PyOx, the ML_2 complexes are clearly terdentate. Comparison of $\beta(ML_2)$ values for oxine (TABLE VI) and for 2-PyOx (TABLE VIII) shows that the latter complexes are more stable* by 3.6 - 4.7 log units, despite a decrease in the basicity (∿0.4 pK_a units) of the donor atoms in 2-PyOx. Furthermore, for the 2-PyOx complexes the usual stability order $\beta(NiL_2) > \beta(CoL_2)$ is observed (TABLE VIII). For oxine ligands containing steric-hindering 2-substituents the reverse order is found(70). By the same reasoning, the ML_2 complexes of 2-HyOx are also terdentate because a non-coordinating, steric-hindering group would have led to reduced stability.

Of the potentially terdentate 2-substituted oxines which have been reported previously(32,34,35,70) only one, 2-aminomethyloxine, forms chelates of a stability comparable to 2-PyOx and 2-HyOx. The similarity of these ligands arises from the flexibility of the 2-substituent in allowing a favorable spatial orientation of the three donor atoms.

Protonated Complexes of 2-HyOx and 2-PyOx

The following section is a discussion of the species with the general formulas $MLHL^+$, MHL^{2+} and $M(HL)_2^{2+}$. These species are formed from the reactions of 2-PyOx and 2-HyOx with the metal ions Co(II), Ni(II),

---

*The 2-PyOx/Cu(II) complex is an exception and is discussed in a later section.

Zn(II) and Cd(II). Particular attention is focused on the species $MLHL^+$. The reactions of Cu(II) with the above ligands are discussed in a later section.

The $MLHL^+$ complexes are particularly stable. For example, if equimolar quantities of Ni(II) and 2-PyOx or 2-HyOx are titrated, the ligand complexes the metal ion mainly in the form $NiLHL^+$, leaving about one-half of the Ni(II) uncomplexed. The stability of $NiLHL^+$ is unexpected, especially since neither $K(NiL_2)$ nor $K(Ni(HL)_2)$ are significantly larger than K(NiL) or K(Ni(HL)) respectively.

The magnitude of the increase in stability for $MLHL^+$ in relation to either ML^+ or MHL^{2+} can be shown by considering the following equations:

$$ML^+ + HL \rightleftarrows MLHL^+, \quad K^{ML}(MLHL)^* = [MLHL^+] / ([ML^+][HL])$$

$$M^{2+} + HL \rightleftarrows MHL^{2+}, \quad K(M(HL)) = [MHL^{2+}] / ([M^{2+}][HL])$$

$$\Delta \log K = \log K^{ML}(MLHL) - \log K(M(HL))$$

It can also be shown that

$$\Delta \log K = \log K^{MHL}(MLHL) - \log K(ML).$$

In the absence of special stabilizing factors $\Delta \log K$ should be < 0. TABLE X shows that log K is positive for all the metal ions except Cu(II).

Neither ML_2 nor $M(HL)_2^{2+}$ exhibit any special stability in relation to ML^+ or MHL^{2+} respectively. However, the presence or absence of a proton in ML_2 or $M(HL)_2^{2+}$ respectively changes the stability of the 1:2 species, $MLHL^+$, in relation to the 1:1 species. Thus, the proton in $MLHL^+$ is either directly or indirectly responsible for the special stability of $MLHL^+$. This

---

*Superscript indicates initial complex.

proton can be associated with one of the oxygen donor atoms for the following three reasons:

Firstly, since the proton in $MLHL^+$ chelates of 2-PyOx dissociates at pH> 5, it is unlikely that the proton is bonded to one of the nitrogen donors because the effect of the coordinated metal ion would be to reduce the pK_a (NH) value for 2-PyOx below 3.62(84)*.

Secondly, for the species MHL^+, the K(M(HL))values for the 2-PyOx system are nearly equal to the corresponding K(ML) values for the 2-,2'-bipyridyl complexes (85), when allowances are made for pK_a(NH) differences. This suggests that at least for the MHL^{2+} complexes of 2-PyOx the metal ion is coordinated to the nitrogen donor atoms and that the proton must be on the oxygen atom. Presumably, the ligand HL in $MLHL^+$ is similarly bonded.

Thirdly, protonated complexes are absent in both 2-PyOx and 2-HyOx/Mn(II) systems. If coordination with both nitrogen donors is required for the formation of $MLHL^+$, then the order of preference which a metal ion exhibits for oxygen and nitrogen could determine whether ML_2 or $MLHL^+$ is formed with a particular metal ion. Since Mn(II) forms more stable complexes with oxygen ligands than with nitrogen ligands(15) it was of interest to examine the complexation data for the 2-PyOx and 2-HyOx/Mn(II) systems to determine if protonated complexes were formed. The formation curve for the two systems (Figure 44 shows the formation curve for the 2-PyOx/Mn(II) system) were typical of oxine-divalent metal-ion systems and there was no indication of protonated complexes. Refinement of the potentiometric data with SCOGS showed that the only predominant

---

*A similar argument also applies to 2-HyOx.

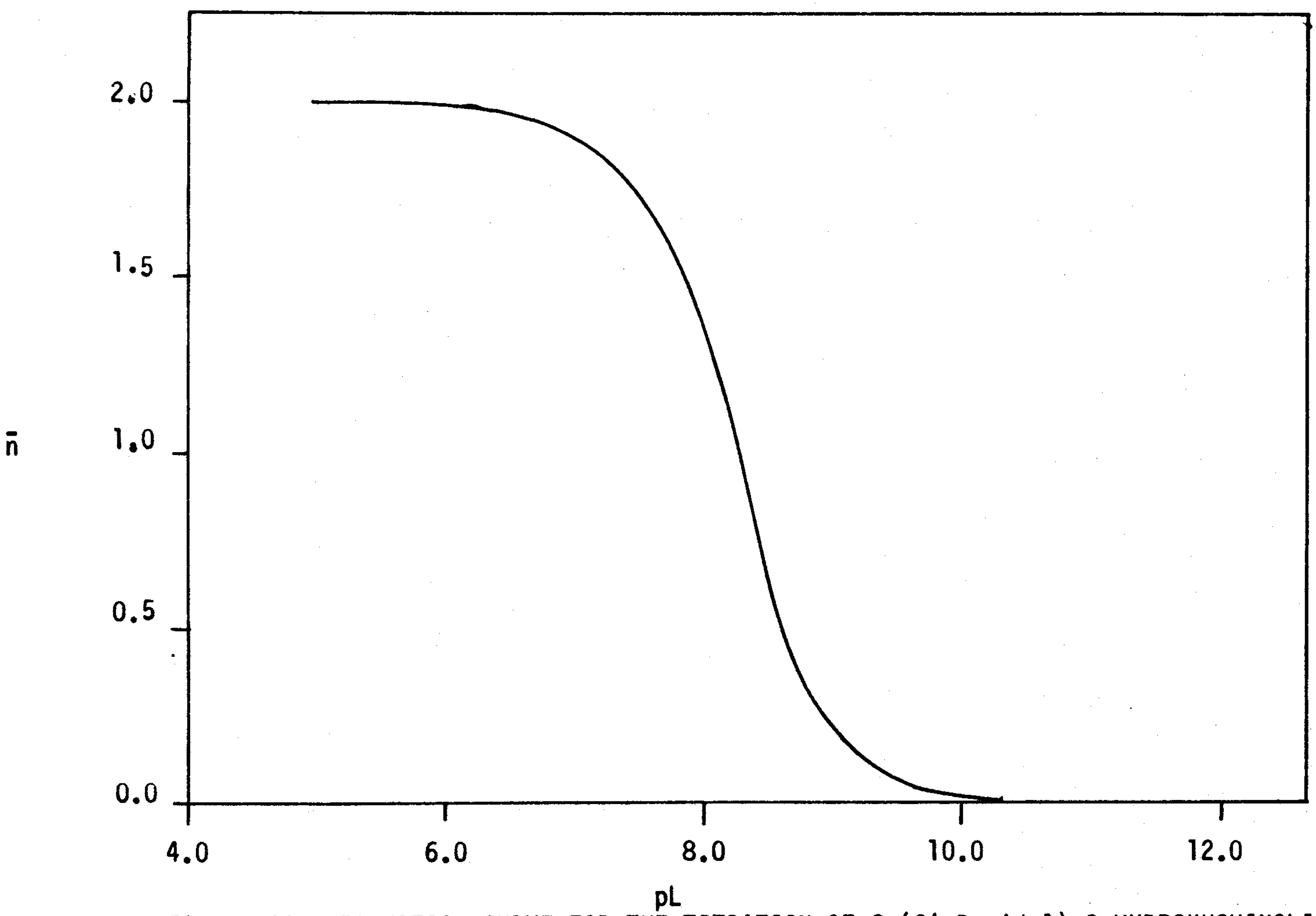


Figure 44. FORMATION CURVE FOR THE TITRATION OF 2-(2'-Pyridyl)-8-HYDROXYQUINOLINE AND MANGANESE(II) IN A 4:1 MOLAR RATIO.

species for both systems were ML^+ and ML_2*. Thus, among other factors, the formation of protonated metal complexes appears to be related to the affinity of the particular metal ion for oxygen and nitrogen donor atoms.

The enhanced stability associated with the addition of the second ligand in $MLHL^+$ cannot be accounted for in terms of spin-pairing (i.e., changes in ground-state configuration). Most probably, the extra stabilization is a result of hydrogen bonding which may occur on addition of the second ligand. In this respect the $MLHL^+$ complexes would resemble the bis (dimethylglyoximato) nickel (II) and other glyoxime complexes, for which a number of structural(86-90) and stability(91-97) studies have been made. In the case of the Ni(II)-dimethylglyoxime complex, solution studies have shown that $K(M(HL)_2) > K(M(HL))$. Although the magnitude of the difference is in doubt, the stabilization of $Ni(HL)_2$ has been attributed to hydrogen bonding(91).

TABLE X shows that the stabilization of $MLHL^+$ is of the order of two log units which represents an energy of ∿ 3 Kcal/mole. Therefore, the degree of stabilization required could easily be provided by hydrogen bonding, the energy of which is 2 - 10 Kcal/mole.

Little direct evidence for hydrogen bonding was obtained in the present work, however. For example, the infrared spectrum (Mull) of $NiLHLClO_4$ (HL = 2-PyOx) showed a broad OH (stretch) band at 3500 cm^{-1} but the position of the band was not indicative of hydrogen bonding. It is possible that this band is due to moisture and that the O-H--O frequency

---

*It is also interesting to note that no protonated species were detected in the 2-PyOx/Al(III) system. Al(III) is a class "A" metal ion like Mn(II).

TABLE X

FORMATION CONSTANTS AND ΔLOG K VALUES FOR 2-PyOx AND 2-HyOx

2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE

Metal Ion	log K^{ML}(MLHL)	log K (MHL)	Δ log K
Co(II)	5.8	4.3	1.5
Ni(II)	7.4	5.4	2.0
Zn(II)	5.4	3.4	2.0
Cd(II)	4.6	3.5	1.1
Ni(II)*	6.4	4.7	1.7
	2-HYDRAZINO-8-HYDROXYQUINOLINE		
Co(II)	6.1	4.3	1.8
Ni(II)	7.3	5.1	2.2
Cu(II)	6.0	9.0	-3.0
Zn(II)	7.2	5.1	2.1
Cd(II)	6.7	5.8	0.9

---

* These formation constants were determined at 60°C.

is shifted to much lower than expected frequencies, as in the Ni(II)/dimethylglyoxime complex(87). In the absence of further work little can be concluded regarding evidence for hydrogen bonding from infrared spectra.

There is, however, indirect evidence which suggests that hydrogen bonding could also be important for the stabilization of $MLHL^+$ in 2-PyOx and 2-HyOx systems. Firstly, within each ligand series the pK_a for $MLHL^+$ (TABLE XI) does not show much variation. Several workers have shown that the pK_a value of an acidic group on a ligand is lowered by complexation and that for a particular series of metal ions the decrease in pK_a is in the order of increasing metal-complex stability(84,99-101). For 2-PyOx and 2-HyOx the stability order for the formation of ML_2 is

$$\mathrm{Ni} > \mathrm{Co} > \mathrm{Zn} > \mathrm{Cd}$$

and

$$\mathrm{Ni} > \mathrm{Co} \sim \mathrm{Zn} > \mathrm{Cd}$$

respectively, whereas the corresponding orders of decreasing acidity are

$$\mathrm{Co} > \mathrm{Zn} > \mathrm{Cd} > \mathrm{Ni}$$

and

$$\mathrm{Co} > \mathrm{Zn} > \mathrm{Ni} > \mathrm{Cd}$$

respectively, which is not what would be expected. It is also interesting to note that the largest difference in stability is 4.4 log units but for the acidity sequence it is only 0.6 log units. Thus, the metal ion is not the sole factor in the determination of the acidity of $MLHL^+$ and the ligands could be important in this respect through the formation of a hydrogen bond between the two oxygen donor atoms in $MLHL^+$ (Figure 45).

Secondly, assuming octahedral coordination, molecular models show that both oxygen donor atoms must be adjacent and that hydrogen bonding between them is possible, as is shown in Figure 45.

TABLE XI

ACID-DISSOCIATION CONSTANTS FOR MLHL CHELATES

Metal Ion	Ligand	pK_a(MLHL)	σ
Co(II)	2-PyOx	5.73	0.04
Ni(II)	2-PyOx	6.37	0.09
Zn(II)	2-PyOx	5.85	0.04
Cd(II)	2-PyPx	6.04	0.04
Co(II)	2-HyOx	7.40	0.06
Ni(II)	2-HyOx	7.76	0.07
Zn(II)	2-HyOx	7.51	0.09
Cd(II)	2-HyOx	7.87	0.07

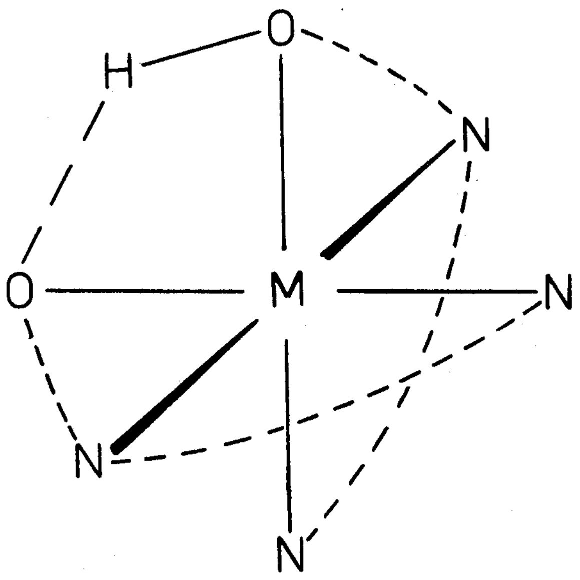


Figure 45. LOCATION OF POTENTIAL HYDROGEN BOND IN $MLHL^+$.

Processes which take place in solution are affected by solvation and entropy changes which may mask changes in formation constants(23,102). Consequently, the formation constants for the 2-PyOx/Ni(II) system were determined at 60.0°C as well as at 25.0°C. Because of the small variations in the formation constants and the inherent inaccuracies of this method for the determination of heat and entropy changes, values of ΔH and ΔS were not calculated. However, TABLE VIII shows that $\beta(NiHL_2)$ does not have an unusually large temperature dependence. The $\Delta \log K$ valuesin TABLE X also show that the stabilization of $NiLHL^+$ is only slightly reduced at 60.0 C. Clarke and Martell(103) have determined the formation constants for the 2,5-diaminopentanoic acid complexes NiL^+, NiL_2 $NiHL^{2+}$, $Ni(HL)_2^{2+}$ and $NiHL^+$. In this system $NiLHL^+$ does not have any special stability so it is unlikely that the corresponding complexes of 2-PyOx and 2-HyOx would be stabilized, relative to ML_2, by solvation due to the presence of a positive charge on $NiLHL^+$. Thus, entropy and solvation forces would seem to be of minor importance in the stabilization of $MLHL^+$, but before a definite conclusion can be made an accurate determination of ΔH and ΔS values is required.

Several examples exist in the literature of protonated metal complexes but many of these involve protonation of an atom remote from the coordination centre of the ligand(103). Ethylenediaminetetraacetic acid forms complexes of the form MHL^{n+}(2) and the Cu(II) complex of N-(8-quinolyl)-salicylamine ($CuLHL^+$)(104) are examples of complexes in which potential donor atoms are protonated. Kumok and Serebrennikov(105) have discussed the stability of protonated complexes in general, but no mention was made of any unusual stability relationships. Johnson and Wilkins(106) have detected the complex $NiHL^{2+}$ for oxine but this complex is rapidly converted into either Ni(II) plus HL or into NiL^+ plus H^+.

The ligand 4,5-dihydroxyacridine forms complexes of the form $M(HL)_2$ but each ligand is bidentate(70). The rigid nature of the ligand introduces excessive strain in the formation of the second chelate ring (per ligand), and thus one of the oxygen donors remains protonated. With the terdentate ligand, 4-amino-5-hydroxyacridine, protonated complexes were not observed(70), perhaps because of the low pK_a(NH) value (2.57)(34) which would result in a destabilization of protonated chelates relative to ML^+ and ML_2. The chelates formed by this ligand were thought to be terdentate but of reduced stability relative to oxine because of chelate ring strain.

The chelates ML^+ and ML_2 for 2-PyOx and 2-HyOx are more stable than the corresponding complexes of 4-amino-5-hydroxyacridine by approximately 5 - 6 log units. Consequently, chelate ring strain would be expected to be much less important for 2-PyOx and 2-HyOx.* However, the looseness of the bite of the three donor atoms of 2-PyOx and 2-HyOx could be partially responsible for the formation of protonated complexes. For example, with 2-PhOx the chelate bite is very tight and protonated complexes are not observed.

In summary, the formation of protonated complexes by a potential terdentate ligand seems to depend on three main factors,

(1) the spatial orientation of the donor atoms or the chelate ring strain,

(2) the availability of two donor atoms which can form a stable bidentate chelate,

---

* Molecular models also show this.

and (3) the nature of the ligand donor atoms in relation to the nature of the metal ion.

The ligand 2-aminomethyloxine is similar to 2-HyOx in both its structure and its acid-base reactions and would seem to fulfill the above three conditions for the formation of protonated complexes. Stevenson and Freiser(35) have reported formation constants for only the ML^+ and ML_2 complexes of 2-aminomethyloxine. The metal ions studied were Co(II), Ni(II), Cu(II), Zn(II) and Mn(II). The failure of the ligand to form protonated species is somewhat puzzling. The following comments show that protonated species could have been overlooked by these workers.

(1) Only one ligand:metal-ion molar ratio (3:1) was used for the potentiometric titrations. The titration curves in this thesis have demonstrated the necessity of using different molar ratios.

(2) In titrations with Ni(II) and Zn(II) precipitation of ML_2 occurred, giving rise to possible misinterpretation of the formation curves. For example, in the 2-HyOx/Ni(II) system in a 2:1 molar ratio (formation curve, Figure 46) precipitation near $\bar{n} = 1$ would cause the formation curve to resemble that expected for step-wise formation of NiL^+ and NiL_2 since the remaining portion of the curve, which is characteristic of the formation of protonated complexes, would not be observed.

(3) Approximation techniques were used to calculate formation constants. Such techniques do not allow detection of protonated species and can lead to inaccurate constants(37,107).

(4) Except for Mn(II) the values of $\beta(ML_2)$ for the Co(II), Ni(II) and Zn(II) chelates of aminomethyloxine are larger than the constants for the corresponding 2-HyOx chelates. This difference between Mn(II) and the

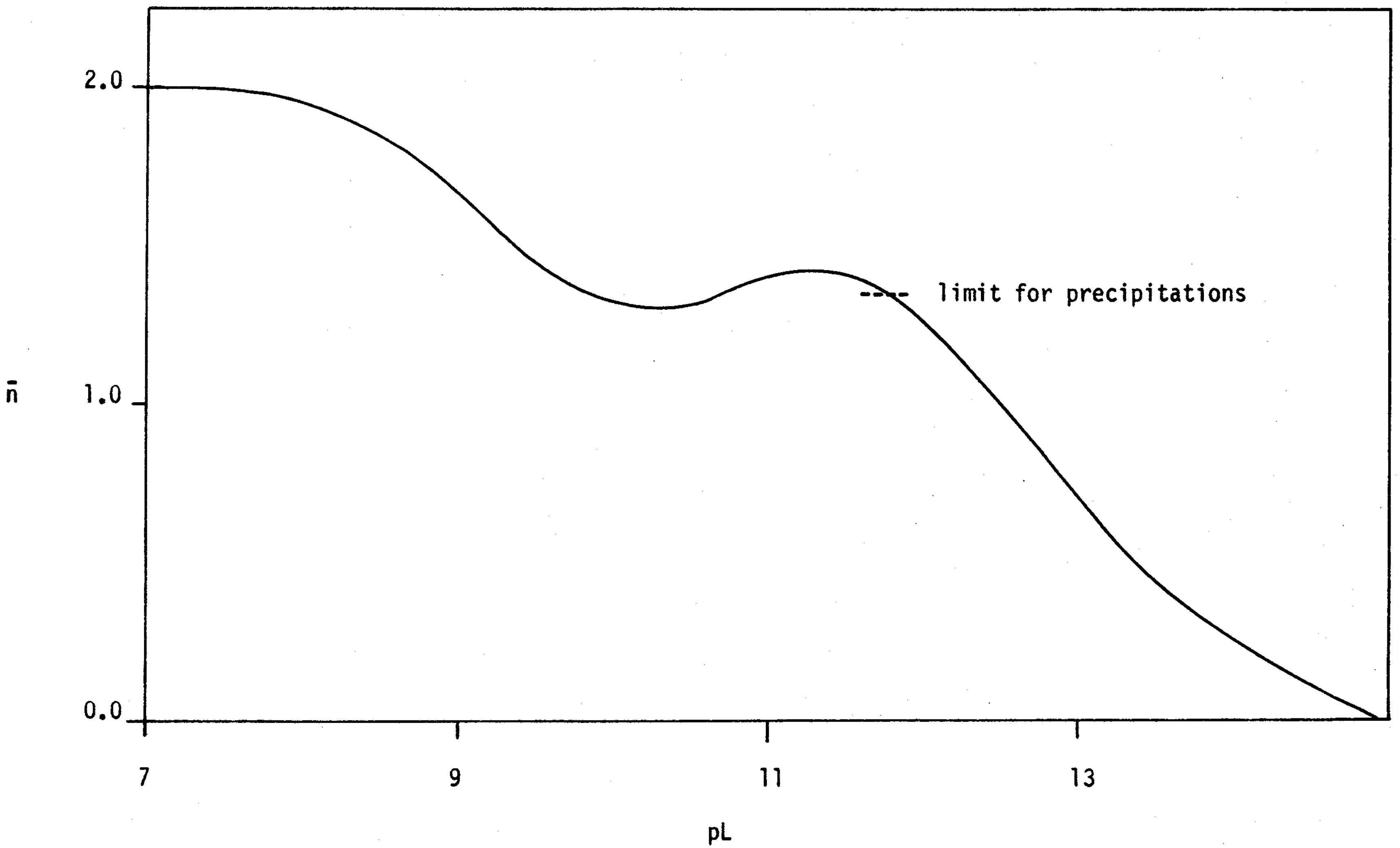


Figure 46. FORMATION CURVE FOR THE TITRATION OF 2-(2'-HYDRAZINO)-8-HYDROXYQUINOLINE AND NICKEL(II) IN A 2:1 MOLAR RATIO.

other metal ions is difficult to explain unless Co(II), Ni(II) and Zn(II) do form protonated complexes. In such a case the liberation of protons would occur prior to the pH range for formation of ML_2. Consequently, the reported values of $\beta(ML_2)$ would be too large. Since Mn(II) does not form protonated complexes with either 2-PyOx or 2-HyOx it is reasonable to assume similar behavior with 2-aminomethyloxine. Therefore, the calculation of formation constants for this chelate is not susceptible to the above error and thus $\beta(MnL_2)$ for 2-HyOx > $\beta(MnL_2)$ for 2-aminomethyloxine.

(5) Stevenson and Freiser explained the formation of CuL_2 on the basis of the following reactions:

$$Cu^{2+} + H_2L^+ \rightleftarrows CuL^+ + 2H^+ \quad (pH \sim 2)$$

$$CuLH_2O^+ \rightleftarrows CuLOH + H^+ \quad (pH \sim 6)$$

$$CuLOH + HL \rightleftarrows CuL_2 + H_2O \quad (pH \sim 9)$$

It was suggested that the last reaction was not manifest in the titration curve since protons are not released or consumed. The following sequence of reactions can also account for the formation of CuL_2.

$$Cu^{2+} + H_2L^+ \rightleftarrows CuL^+ + 2H^+ \quad (pH \sim 2)$$

$$CuL^+ + H_2L^+ \rightleftarrows CuLHL^+ + H^+ \quad (pH \sim 6)$$

$$CuLHL^+ \rightleftarrows CuL_2 + H^+ \quad (pH \sim 8)$$

Since the last reaction would occur in the pH range of the acid-dissociation of H_2L^+ ($pK_a(NH) = 8.56$), it would be obscured and readily confused with that equilibrium.

It would be of interest to reinvestigate the 2-aminomethyloxine chelate systems and other ligands such as 2-(2'-aminophenyl)oxine and 2-aminoethyloxine.

Reactions of 2-PyOx and 2-HyOx with Cu(II)

The reactions of Cu(II) with 2-PyOx and 2-HyOx are somewhat different from those of Co(II), Ni(II), Zn(II) and Cd(II) and are discussed below.

The three main differences are:

(1) the values of $K(CuL)/K(CuL_2)$ are unusually large (TABLES VIII and IX);

(2) the complex $CuLHL^+$ is not stabilized, as is shown by the formation of mainly CuL^+ at a 1:1 molar ratio and by the negative value of Δlog K for the 2-HyOx chelate (TABLE X),

and (3) the species $CuHL^{2+}$ and $Cu(HL)_2^{2+}$ are not formed in the PyOx system.

Almost all hexacoordinate Cu(II) compounds for which crystal structures have been determined are tetragonally distorted(108). This behavior has been interpreted in terms of the Jahn-Teller effect(108) but coulombic forces, which have nothing to do with this effect, can also be important(109). The tetragonal distortion of Cu(II) complexes is well documented in both the solid and liquid state and the net result is a stabilization of the four planar bonds and a weakening of the two axial bonds. For example, ethylenediamine (en) forms mainly $Cuen^{2+}$ and $Cu(en)_2^{2+}$ (110) as is shown by the formation constants in TABLE XII. $Cu(en)_2^{2+}$ is considered to be a distorted octahedron with four coplanar nitrogen atoms

TABLE XII

FORMATION CONSTANTS OF SOME COPPER (II) CHELATES(111)

Ligand	pK_a (NH)	log K (ML)	log K (ML)	log K (ML)
2,2'-bipyridyl	4.40	8.1	5.5	3.4
1,10-phenanthroline	4.98	9.0	6.7	5.0
ethylenediamine	7.3 , 9.8	10.7	9.3	1.0

and two water molecules in the axial positions(112). For 1,10-phenanthroline and 2,2' - bipyridyl, however, the second ligand is <u>cis</u> to the first(113). Since bond formation in the plane is favored relative to the axial position, K(CuL_2) is much lower than K(CuL) for these two ligands in relation to the difference between K($Cu(en)_2$) and K(Cuen) (TABLE XII).

If the Cu(II) complexes of 2-PyOx and 2-HyOx are also tetragonally distorted, CuL^+ should be stabilized relative to CuL_2 since the first ligand forms three bonds in the plane as compared to only one for the second ligand. The large ratio of K(CuL) / K(CuL_2) for the 2-PyOx and 2-HyOx complexes could then be explained. The stabilization could also account for the observation that CuL^+ is formed in preference to $CuLHL^+$ when the molar ratio is 1:1. It also follows that CuL^+ and $CuLHL^+$ should predominate over $CuHL^{2+}$ and $Cu(HL)_2^{2+}$ at molar ratios > 1:1.

Finally, the fact that 2-HyOx forms observable amounts of $CuHL^{2+}$ and $Cu(HL)_2^{2+}$ and 2-PyOx does not can likely be attributed to the larger pK_a (NH) value for 2-HyOx which would help stabilize the 2-HyOx complexes relative to those of 2-PyOx.

Discussion of Errors

Many of the results in this work were obtained from the analysis of potentiometric titration data. The factors contributing to the accuracy of these analyses have been discussed elsewhere(114,115). In summary, the main errors arise from (1), the measurement of pH and (2), the interpretation of the potentiometric data in terms of the number and kind of species present. The first error arises essentially from the uncertainty in the pH values of the standard buffer solution used for pH-meter calibration. These errors affect the values of the formation constants and can be minimized by using the same conditions for a series of titrations. This technique permits the comparison of protonation and formation constants with some certainty.

The second error can be minimized in two ways. Firstly, if titrations are performed at different ligand and metal-ion concentrations. In this work, as many as eight and as few as three sets of concentrations were titrated for each system. Secondly, if a number of auxillary techniques are used to confirm the number and kind of species present. In this study the interpretation of the titration curves was confirmed by the formation curves, α_A plots, ultraviolet-visible spectra, mole ratio plots, polarography, i.r. spectra and elemental analysis.

Suggestions for Further Work

(1) A more detailed i.r. investigation of the 2-PyOx and 2-HyOx complexes $MLHLClO_4$ is required to obtain evidence for hydrogen bonding. For example, the product of hydrogen/deuterium exchange reactions should be examined by i.r.

(2) The reactions of a number of oxines towards Al(III) should be investigated in an attempt to definitely show that the slow reactions of polynuclear-hydroxo Al(III) species are an important factor in the apparent lack of reactivity of many 2-substituted oxines with Al(III) in aqueous solution.

(3) In view of the possible misinterpretation of the 2-aminomethyl-

oxine systems, this ligand should be reinvestigated using the techniques in this work. It would also be of interest to prepare and study the reactions of 2-aminoethyloxine and 2-(2'-aminophenyl)oxine.

SUMMARY

(1) The new compound, 2-(2'-pyridyl)-8-hydroxyquinoline, and 2-(2'-hydroxyphenyl)-8-hydroxyquinoline and 2-hydrazino-8-hydroxyquinoline have been prepared as ligands and characterized by proton n.m.r., mass spectroscopy, i.r. spectroscopy and elemental analysis.

(2) The acid-base reactions of the above ligands have been studied by spectroscopic and potentiometric titration techniques, and acid-dissociation constants have been calculated from the potentiometric data.

(3) Qualitative tests showed that the selectivity of the ligands is not high and that differences among the ligands in their reactions towards metal ions is consistent with differences in hard and soft character of the donor atoms.

(4) With divalent metal ions 2-(2'-hydroxyphenyl)-8-hydroxyquinoline forms 1:1 terdentate complexes of the general formula, ML. With Al(III), this ligand forms AlL^+ and AlL_2 and is the first 2-substituted 8-hydroxyquinoline known to react appreciably with Al(III) in acid solution in the absence of a secondary ligand. This difference in behavior towards Al(III) has been discussed.

(5) With Co(II), Ni(II), Zn(II) and Cd(II), 2-(2'-hydroxyphenyl)-8-hydroxyquinoline and 2-(2'-pyridyl)-8-hydroxyquinoline form the non-protonated complexes ML^+ and ML_2 and the protonated species $MLHL^+$, MHL^{2+} and $M(HL)_2^{2+}$, in which the proton is located on the quinolinic oxygen atom. The species $MLHL^+$ is stabilized relative to the 1:1 species, likely because of hydrogen bonding. With Mn(II), only the species ML^+ and ML_2 are formed. The absence

of protonated species is probably the result of the greater affinity of Mn(II) for oxygen, compared to the other divalent ions studied. The reactions of Cu(II) also differ from those for the above metal ions and these differences have been interpreted in terms of tetragonal distortion in the Cu(II) complexes.

(6) The factors which affect the formation of protonated complexes have been discussed.

(7) The data reported previously for 2-aminomethyl-oxine has been examined. It is felt that the reactions of this ligand should be reinvestigated.

(8) The program SCOGS has been revised and an error, which was present in the original SCOGS, has been corrected.

(9) A new interpretation of formation curves in terms of the number of liberated protons per metal ion has been described.

APPENDIX I

Synthesis of Intermediate Compounds and Ligands

8-Methoxyquinoline. This compound was prepared by a modified Skraup reaction as described by King and Sherred(116). Vacuum distillation (1 - 5 mm Hg) at 110 - 120°C gave a clear pale-yellow liquid which solidified on standing at room temperature. The yield varied from 40 - 70 percent. Maximum yields were obtained when rapid elevations in temperature were avoided. Appreciable charring occurred when the temperature rose quickly.

n - Butyllithium. The procedure of Gilman *et al.*(117) was followed and the molarity of the resulting solution was determined as described by Gilman and Haubein(118). The yield varied from 70 to 85 percent.

2 - Pyridyllithium. This compound was prepared via a halogen-metal interconversion reaction according to the procedure of Gilman and Spatz(119). The temperature of the reaction mixture was maintained at -40°C as suggested by Klingsberd(120) since this temperature resulted in higher yields. The final product was used within 20 minutes.

2-(2'- Pyridyl)-8-hydroxyquinoline. The following procedure is similar to that for the preparation of many 2-substituted oxines (51,52) and 2-substituted quinolines(54). Twenty-five grams (0.2 moles) of 8-methoxyquinoline, dissolved in a minimum amount of anhydrous diethyl ether, was added dropwise over a period of 20 to 30 minutes to a solution of 2-pyridyllithium, prepared from 38 g (0.25 moles) of 2-bromopyridine as referred to above. During this addition the reaction mixture was stirred vigorously and the temperature was controlled at -40°C. An atmosphere of nitrogen was maintained in the flask throughout the reaction. Stirring was continued for 2½ hours and the mixture was allowed to slowly reach room temperature. The mixture, consisting of an orange solution and an orange-yellow precipitate, was poured over ice and the aqueous layer was extracted several times with ether. The ether extracts were evaporated under vacuum (water pump) to yield a red residue which was refluxed in 50 ml of nitrobenzene for 30 minutes. The resulting solution was distilled under vacuum (1 - 5 mm Hg) and 2-(2'- pyridyl)-8-methoxyquinoline, a viscous yellow oil, was collected over the temperature range 180 - 200°C. Five grams of this compound were refluxed with 50 ml of 48% hydrobromic acid for 24 hours. The mixture was then cooled to 0°C, filtered and the precipitated hydrobromide salt was dissolved in water. The resulting solution was made strongly basic with 1 N sodium hydroxide, filtered to remove undissolved material, and then neutralized to yield a light cream-colored precipitate. The insoluble residue was treated with 200-ml portions of 1 N sodium hydroxide until no further precipitation occurred on neutralization. Repeated recrystallization from aqueous ethanol gave pure 2-(2'-pyridyl)-8-hydroxyqinoline. M.p., 129-130°C; yield, 9 g (20% based on 8-methoxyquinoline).

Calculated for $C_{14}H_{10}ON_2$: C, 75.6%; H, 4.51%; N, 12.6%.
Found: C, 75.5%; H, 4.44%; N, 12.5%.

<u>2-Lithiumlithiumphenoxide</u>. This compound was prepared via a halogen-metal interconversion reaction between 2-bromophenol and n-butyllithium, as described by Gilman and Arntzen[(54)]. This compound was used within

10 minutes.

2-(2'-hydroxyphenyl)-8-methoxyquinoline. The following synthesis is similar to that used by Towle(55). To a solution of 2 lithiumlithiumphenoxide prepared from 14.2 g (0.082 moles) of 2-bromophenol, 12.0 g (0.075 moles) of 8-methoxyquinoline in 100 ml of anhydrous ether were added dropwise over a 30-minute period with constant stirring. The orange reaction mixture was stirred for an additional 2 hours and then poured over ice. The water layer was removed and neutralized to yield a mixture of a solid and an oil. Complete recovery of the product was ensured by extraction of the ether layer with 200 ml-aliquots of 2 N sodium hydroxide until no further precipitation occurred on neutralization of the basic solution. The product, a dihydro compound, was partially purified by recrystallization from aqueous ethanol. A flaky yellow solid was obtained which melted at 91 - 95°C. Two grams of this compound were dissolved in 50 ml of nitrobenzene and refluxed for one hour. On cooling, this solution was extracted with 200 ml of 2 N hydrochloric acid, which yielded a yellow solid on neutralization. Recrystallization from aqueous ethanol produced bright yellow crystals of 2-(2'-hydroxyphenyl)-8-methoxyquinoline. M.p., 160 - 161°C (literature, 162 - 164°C(55)); yield, 4 g (20% based on 8-methoxyquinoline).

2-(2'-hydroxyphenyl)-8-hydroxyquinoline. Four grams (0.016 moles) of 2-(2'-hydroxyphenyl)-8-methoxyquinoline were placed in 50 ml of 48% hydrobromic acid and refluxed for 24 hours. On cooling to 0°C, the precipitated hydrobromide salt was filtered off and dissolved in water. This solution was made basic and filtered to remove any insoluble residue. The filtrate was neutralized and the orange precipitate was recrystallized from aqueous ethanol to give pure 2-(2'-hydroxyphenyl)-8-hydroxyquinoline. M.p., 116 - 117°C(literature, 117 - 118°C(54)); yield, 3 g (80%)

Calculated for $C_{15}H_{11}O_2N$: C, 75.9%; H, 5.06%; N, 6.05%.
Found: C, 75.9%; H, 4.78%; N, 5.91%.

8-Hydroxyquinoline N-oxide. This compound was prepared by oxidizing oxine with peracetic acid, as described by Phillips, Barral and Breese(121). Recrystallization from aqueous ethanol yielded a yellow-brown solid. M.p., 136 - 137°C (literature, 137°C(121)); yield, 37 g (70%).

2-Hydroxy-8-acetoxyquinoline. Thirty-seven grams of 8-hydroxyquinoline N-oxide were mixed with sufficient acetic anhydride to allow free movement of a magnetic stirrer, and then the mixture was stirred for three days. After one day, the solid dissolved to give a clear red-brown solution. After an additional day, a white solid appeared. Water was added to remove the excess acetic anhydride and on neutralization with 1 N ammonia a white solid was obtained. M.p., 242 - 245°C (literature, 250°C (121)); yield, 33 g (80%).

2,8-dihydroxyquinoline. The procedure of Phillips, Barral and Breese (121) was followed but the product obtained had a significantly higher melting point. After recrystallization from aqueous ethanol, the product melted at 287 - 289°C (literature, 250°C(121)). The yield was 20 g (72%).

2-hydroxy-8-tosyloxyquinoline. Twenty grams (0.12 moles) of 2,8-dihydroxyquinoline in 250 ml of water containing 40 g of sodium carbonate were added to 40 g (0.12 moles) of p-toluenesulphonylchloride in 200 ml of chloroform. This mixture was stirred vigorously for 24 hours and the chloroform layer was then separated and evaporated (water pump). Purification of the remaining residue was extremely difficult. After several recrystallizations from aqueous ethanol, the compound was judged to be of sufficient purety for synthetic purposes. M.p., 203 - 207°C (literature, 210 - 211°C(122)); yield, 25 g (65%).

2-chloro-8-tosyloxyquinoline. Twenty-five grams (0.08 moles) of 2-hydroxy-8-tosyloxyquinoline were mixed with 72 g (0.35 moles) of phosphorous pentachloride, to which a small amount of phosphorous oxychloride had been added. This mixture was then treated according to procedure of Hamana *et al*. M.p., 129 - 131°C (literature, 132 - 134°C(122)); yield, 21 g (80%).

2-chloro-8-hydroxyquinoline. This compound was prepared as described by Rudolph *et al*.(58). The p.m.r. spectrum was identical to that given elsewhere(59). M.p., 82 - 83°C (literature, 83 - 84°C (58)); yield 90%.

2-hydrazino-8-hydroxyquinoline. This compound was prepared as described by Rudolph *et al*.(58). M.p., 177 - 178°C (literature, 177 - 178°C (58)); yield, 53%.

Calculated for $C_9H_{10}ON_3$: C, 62.2%; H, 5.70%; N, 24.0%

Found: C, 61.8%; H, 5.51%; N, 24.0%.

APPENDIX II

DERIVATION OF EQUATIONS

Spectrophotometric Determination of Acid-Dissociation Constants

The final acid-dissociation (phenolic-OH group) of 2-PhOx can be represented by

$$HL^{-} \rightleftarrows H^{+} + L^{=}, \; K_a = [H^{+}][L^{=}]/[HL^{-}]$$

The absorbence As, of a solution containing $L^{=}$ and HL^{-} is given by

$$As = \Sigma_0 [HL^{-}] + \Sigma_1 [L^{=}] \qquad [1]$$

where Σ_0 and Σ_1 are molar absorptivities for HL^{-} and $L^{=}$, respectively, when 1-cm cells are used. The mass balance is

$$C_L = [HL^{-}] + [L^{=}] \qquad [\,2\,]$$

assuming the concentrations of other forms of the ligand to be negligible. From the expression for K_a and [2], the following expressions are obtained:

$$[L^{=}] = C_L - [H^{+}]\, C_L/(K_a + [H^{+}]) \qquad [3]$$

and

$$[HL^{-}] = [H^{+}]\, C_L/(K_a + [H^{+}]) \qquad [4]$$

From [1],[3], and [4],

$$As = \Sigma_0\, C_L\, [H^{+}]/(K_a + [H^{+}]) + \Sigma_1\, C_L\, K_a/(K_a + [H^{+}]) \qquad [5]$$

Equation [5] has two unknowns, K_a and Σ_1. Σ_0 is known from the ultraviolet-visible spectra of 2-PhOx at pH 12. The use of data at two different points, $\{As_1, [H^{+}]_1\}$ and $\{As_2, [H^{+}]_2\}$ allows calculation of K_a from

$$K_a = \{\Sigma_0\, C_L\, ([H^{+}]_2 - [H^{+}]_1) + As_1 [H^{+}]_1 - As_2 [H^{+}]_2\} / (As_2 - As_1) \qquad [6]$$

The data for this study is listed in APPENDIX III.

Potentiometric Determination of Acid-Dissociation (Concentration) Constants

The derivation below is for the calculation of the pK_a (NH) and pK_a(OH) values for 2-PyOx, 2-HyOx and oxine. The equations for 2-PhOx are similar. The fully protonated form of 2-PyOx and 2-HyOx can be represented by H_3L^{2+}. Since this species is a very strong acid and is not observed even in the most concentrated acid solutions used (∿ pH 2), it has been ignored in the following treatment.

$$H_2L^+ \rightleftarrows H^+ + HL \text{ , } K_a(NH) = [H^+][HL]/[H_2L^+]$$

$[ClO_4^-]$ = concentration of strong acid ($HClO_4$)

C_L = total concentration of weak acid

$$= [H_2L^+] + [HL] + [L^-] \qquad [7]$$

$$K_w = [H^+][OH^-] \qquad [8]$$

The charge balance is

$$[Na^+] + [H^+] + [H_2A^+] = [OH^-] + [ClO_4^-] \qquad [9]$$

From [8] and [9]

$$[H_2A] = K_w/[H^+] + [ClO_4^-] - [H^+] - [Na^+] \qquad [10]$$

From [7] and [10]

$$[HA] = C_A - K_w/[H^+] - [ClO_4^-] + [H^+] + [Na^+] \qquad [11]$$

From [10] and [11] and the pH-meter readings, the value of K_a (NH) can be calculated for each point in the titration. The derivation of equations for the acid-dissociation of HL is similar. The values of pH chosen for the equations [10] and [11] were limited to $\bar{p}$ values between 1.2 - 1.8, as explained under EXPERIMENTAL. Values of $\bar{p}$ can be obtained from [10] and [11] and

$$\bar{p} = (2[H_2L^+] + [HL]) / ([H_2L^+] + [HL]) \qquad [12]$$

The data for these studies is listed in APPENDIX III.

Potentiometric Determination of Formation Constants

Oxine/Divalent Metal-Ion Systems. The formation constants are defined by

$$K(ML) = [ML^+] / ([M^{2+}][L^-]) \text{ and } K(ML_2) = [ML_2] / ([ML^+][L^-])$$

The metal ion and ligand mass-balance equations are

$$C_M = [M^{2+}] + [ML^+] + [ML_2] \qquad [13]$$

$$C_L = [ML^+] + 2[ML_2] + [H_2L^+] + [HL] + [L^-] \qquad [14]$$

The charge balance is

$$[Na^+] + [H^+] + [H_2L^+] + 2[M^{2+}] + [ML^+] = [ClO_4^-] + [OH^-] + [L^-] \qquad [15]$$

where $[ClO_4^-]$ is the total perchlorate concentration from metal perchlorate and perchloric acid. From the acid-dissociation constant expressions for the ligand and rearrangement of [13] and [14] in terms of $[M^{2+}]$ and $[ML_2]$, the following expression is obtained for $[L^-]$ by substitution into [15]

$$L^- = ([ClO_4^-] + [OH^-] - [Na^+] - [H^+] - 2C_M + C_L) /$$
$$([H^+] / K_a(OH) + 2[H^+]^2 / (K_a(OH)\, K_a(NH))) \qquad [16]$$

From [16], $[L^-]$ can be calculated for each point in the titration.

The parameter, $\bar{n}$, is given by

$$\bar{n} = ([ML^+] + 2[ML_2])/C_M \qquad [17]$$

From [14] and [17]

$$\bar{n} = (C_L - [L] - [HL] - [H_2L^+])/C_M \qquad [18]$$

From [16] and the expressions for the acid-dissociation constants of the ligand, the value of $\bar{n}$ can be calculated for each point in the titration. From [17] and the expressions for the formation constants the following expression is obtained:

$$\bar{n} = (K(ML)\,[L^-] + 2K(ML)\,K(ML_2)\,[L^-]^2)/(1 + K(ML)\,[L^-] + K(ML)\,K(ML_2)\,[L^-]^2) \qquad [19]$$

On rearrangement, [19] yields

$$\bar{n}/\,[(1-\bar{n})\,[L^-]] = K(ML)\,K(ML_2)\,(2-\bar{n})\,[L^-]\,/(1-\bar{n}) + K(ML_2) \qquad [20]$$

This equation is of the form

$$Y = K(ML)\;\;K(ML_2)\,X + K(ML) \qquad [21]$$

which is an equation for a straight line with slope = $K(ML)\ K(ML_2)$ and intercept = $K(ML)$. Linear least squares refinement gives values of $K(ML)$ and $K(ML_2)$ which can be further refined by minimizing the following expression,

$$\sum_i [\bar{n}_i - (K(ML)[L^-]_i + 2K(ML)\,K(ML_2)[L^-]_i^2)\,/\,(1 + K(ML)[L^-]_i + K(ML)\,K(ML_2)[L^-]_i) \qquad [22]$$

by means of an iterative procedure in which the values of $K(ML)$ and $K(ML_2)$ are varied.

Oxine/Trivalent Metal-Ion Systems. Equations for $[L^-]$ and $\bar{n}$ can be obtained in a fashion similar to those obtained above. Approximate values of the formation constants can be obtained by the following method:

at $\bar{n} = 1.0$, [19] gives

$$1 = K(ML)\ K(ML_2)\ [L^-]^2 + 2\ K(ML)\ K(ML_2)\ K(ML_3)\ [L^-]^3 \qquad [23]$$

which reduces to

$$K(ML)\ K(ML_2) = 1/(\ [L^-]^2 (1 + 2\ K(ML_3)\ [L^-]\ )) \qquad [24]$$

If $2\ K(ML_3) < 1$, then $K(ML)\ K(ML_2) \approx 1/[L^-]^2$ [25]

Similarly, at $\bar{n} = 0.5$, if $5\ K(ML)\ K(ML_2)\ K(ML_3)\ [L^-]^3$ is small

then $1 = K(ML)\ [L^-] + 3\ K(ML)\ K(ML_2)\ [L^-]^2$ [26]

Let $[L^-]$ at $\bar{n} = 0.5$ and $n = 1.0$ be $[L^-]_{0.5}$ and $[L^-]_1$ respectively, then

substitution of [25] into [26] leads to the following expression for $K(ML)$:

$$K(ML) = 1/[L^-]_{0.5} - 3\,[L^-]_{0.5}/[L^-]_1^2 \qquad [27]$$

Substitution of [27] into [25] or [26] gives an expression for $K(ML_2)$. $K(ML_3)$ can be obtained in a similar manner at $\bar{n} = 2.0$. These approximate values are refined by minimizing the following expression with respect to $K(ML_2)$ and $K(ML_3)$:

$$\sum_i [\,\bar{n}_i - (K(ML)\,[L^-]_i + 2\ K(ML)\ K(ML_2)\,[L^-]_i^2 + 3\ K(ML)\ K(ML_2)\ K(ML_3)\,[L^-]_i^3\ )\ /\ (1 + K(ML)\,[L^-]_i + K(ML)\ K(ML_2)\,[L^-]_i^2 + K(ML)\ K(ML_2)\ K(ML_3)\,[L^-]_i^3\ )\,]^2 \qquad [28]$$

2-PhOx/Divalent Metal-Ion Systems. The procedure for the derivation of equations for 2-PhOx metal chelates is similar to that for the oxine metal chelates. Since there is one less unknown for 2-PhOx, a graphical procedure is not required. The expressions for $[L^{=}]$, $\bar{n}$, and K(ML) are

$$[L^{=}] = ([ClO_4^-] + [OH^-] + 2C_L - [Na^+] - [H^+] - 2[M^{2+}]) / ([H^+] / K_a(OH_2) + 2[H^+]^2/(K_a(OH_2)\ K_a(OH_1))\ ^* + 3[H^+]^3 / (K_a(OH_2)\ \ K_a(OH_1)\ \ K_a(NH)))] \qquad [29]$$

$$\bar{n} = C_L - [L^{=}]([H^+]/K_a(OH_2) + [H^+]^2/ (K_a(OH_2)\ K_a(OH_1)) + [H^+]^3/ (K_a(OH_2) = K_a(OH_1)\ K_a(NH))) \qquad [30]$$

and $$K(ML) = \bar{n}/ [(1 - \bar{n})\ [L^{=}]] \qquad [31]$$

2-HyOx and 2-PyOx/Divalent Metal-Ion Systems. Before the program SCOGS could be used for the calculation of formation constants for these metal complex systems, a number of program changes were necessary. Some, such as conversion from Fortran IV (E) to Fortran IV, involved extensive deck alterations but most changes were minor modifications of input and output.

After some experience with the program, it was noticed that the values obtained for the constants depended on how NDP (number of dissociable protons on the ligand in the form (e.g., H_2L^+,HL or L^-) it was added) was read into the computer. For example, if the ligand was added as NaL (NDP = o) rather than as HL (NDP = 1), while the potential acidity was kept constant by

---

*OH_1 and OH_2 refer to the quinolinic and phenolic OH groups, respectively.

alteration of the strong acid concentration, significant variations in the formation constants were observed (TABLE I). This behavior indicated an error in the program.

The final values of the formation constants are ultimately determined by the degree of agreement between a calculated titre (i.e., volume) of base (CTITR) and the actual observed titre, since the least squares procedure in SCOGS is based on the difference between these two terms. From the input constants and the initial concentrations of acid, ligand and metal ion, the concentrations of all the species in the solution are calculated. From the calculated concentrations, a value of CTITR can be obtained as follows:

$$\mathrm{CTITR} = [\, A + D - (\, X + F\,(\beta_i))\,] \,/\, \mathrm{BA} \qquad [32]$$

where A = total available moles of H^+ from the ligand,

D = total moles of H^+ from the addition of strong acid ($HClO_4$),

X = moles of free H^+ (i.e., the equilibrium concentration at each point in the titration) as measured by pH-meter,

$F(\beta_i)$ = moles H^+ still bound by various species at each point in the titration (the concentration of these species are a function of the overall formation constants β_i),

and BA = concentration of the standard base.

The values of A and X are known experimentally and should be used as such. In the original program, however, this procedure was not followed. This is clearly incorrect since it is the deviation of CTITR from the actual titre caused by the use of <u>only</u> approximate β_i values that is of interest.

The SCOGS program was corrected as outlined below. The equation for

TABLE I

EFFECT OF NDP ON FORMATION CONSTANTS

(2-PyOx / Mn(II) System)

Program	log K(ML)	log K(ML_2)	log β(ML_2)
Modified Rossotti and Rossoti Least Squares	8.444	8.051	16.495
SCOGS (NDP = 1)	8.399	8.168	16.567
SCOGS (NDP = 0)	8.165	7.930	16.095
Revised SCOGS (NDP = 1)	8.443	8.055	16.498
Revised SCOGS (NDP = 0)	8.444	8.055	16.499

CTITR in the original program(40) (line 158) is as follows:

CTITR = V(K) * (TURM(K) + AC (K) - HO)/(BA - TURM(K) + HO) [33]

where TURM(K) = concentration (corrected for volume changes) of available protons from the ligand at point K in the titration (the value of TURM(K) depends on the value of NDP),

V(K) = initial volume of the solution, and is always the value for K = 1,

AC(K) = initial strong-acid concentration, and is always the value for K = 1,

and HO = concentration of protons bound by various species, as calculated with β_i, plus the concentration of free protons as measured by the pH-meter.

Equation [33] was derived by solution of the following expression for CTITR(123):

CTITR = [TURM(K) * (V(K) + CTITR) + V(K) * AC(K) - HO * (V(K) + CTITR)]/BA [34]

The volume (V(K) + CTITR) in [34] is used to convert the concentration terms, TURM(K) and HO, into values of A, X and F(β_i) in [32] . Since the value of actual titre ≠ CTITR, the values of A, X and F (β_i) are not equivalent to the known experimental values. The final values of β_i are refined by the minimization of $\sum_K$ (actual titre - CTITR$)^2_K$ and thus the final values of β_i will be in error. In regards to term A, the value of TURM(K) depends on NDP and thus the magnitude of the error in V(K) + CTITR will change with a change in the value of NDP. The correct equation for

CTITR should be as follows:

CTITR = [TURM(K) * VOX1(K) * AC(K) + HO * VOX1(K)] /BA [35]

where VOX1(K) = actual volume at point K in the titration.

APPENDIX III

SPECTROPHOTOMETRIC DATA

Determination of pK_a(OH) for Phenolic-OH Group of 2-PhOX

TABLE I

λ = 302 nm

As_1	As_2	$[H_1^+] \times 10^{+14}$	$[H_2^+] \times 10^{+14}$	$K_a \times 10^{+15}$
0.260	0.175	9.36	2.75	4.5
0.260	0.180	9.36	3.34	3.5
0.260	0.156	9.36	2.34	3.8
0.221	0.180	5.85	3.34	4.6
0.221	0.156	5.81	2.34	4.6
0.195	0.156	4.25	2.34	5.3
0.195	0.180	4.25	3.34	6.2

λ = 291 nm

As_1	As_2	$[H_1^+] \times 10^{+14}$	$[H_2^+] \times 10^{+14}$	$K_a \times 10^{+15}$
0.302	0.251	5.85	3.34	4.8
0.302	0.226	5.85	2.34	6.0
0.350	0.226	9.36	2.34	5.2
0.350	0.248	9.36	2.75	5.8
0.350	0.276	9.36	4.25	5.0
0.276	0.226	4.25	2.34	5.6

$K_a = 5.4 \pm 0.6 \times 10^{-15}$

POTENTIOMETRIC TITRATION DATA

Representative Data for Acid-Dissociation Constants

TABLE II

8-HYDROXYQUINOLINE

(initial volume = 100 ml, 25.0°C, [$HClO_4$] = 0.00986, [NaOH] = 0.1254, 0.0630 g reagent, 50% (v/v) aqueous dioxane)

pH	VOLB[*]	$\bar{p}$	pK_a
3.612	1.0	1.783	4.100
3.722	1.2	1.740	4.105
3.826	1.4	1.693	4.109
3.922	1.6	1.643	4.107
4.020	1.8	1.592	4.111
4.158	2.1	1.512	4.109
4.305	2.4	1.431	4.114
4.408	2.6	1.376	4.117
4.578	2.9	1.293	4.125
4.780	3.2	1.209	4.131

pK_a(NH) = 4.11 (mean)

Standard Deviation = 0.01

pH	VOLB[*]	$\bar{p}$	pK_a
10.612	4.8	0.756	11.103
10.789	5.1	0.672	11.100
10.940	5.4	0.588	11.095
11.078	5.7	0.504	11.086
11.211	6.0	0.423	11.078
11.496	6.5	0.296	11.123

pK_a(OH) = 11.09 (mean)

Standard Deviation = 0.02

---

[*] VOLB refers to volume of base in ml. The volumes are significant to three places past the decimal but the zeros have been omitted.

TABLE III

2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE

(initial volume = 100 ml), 25.0°C, $[HClO_4] = 0.00986$, $[NaOH] = 0.1254$, 0.0999 g reagent, 50% (v/v) aqueous dioxane)

pH	VOLB	$\bar{p}$	pK_a
3.072	0.3	1.800	3.606
3.139	0.5	1.774	3.607
3.208	0.7	1.745	3.607
3.282	0.9	1.714	3.612
3.355	1.1	1.679	3.612
3.428	1.3	1.640	3.611
3.502	1.5	1.599	3.610
3.578	1.7	1.560	3.610
3.658	1.9	1.512	3.612
3.738	2.1	1.465	3.611
3.828	2.3	1.418	3.618
3.920	2.5	1.370	3.623
4.020	2.7	1.321	3.628
4.132	2.9	1.271	3.635
4.258	3.1	1.219	3.641
		pK_a(NH) =	3.61 (mean)
		Standard Deviation =	0.01
10.798	4.8	0.766	11.248
10.917	5.0	0.712	11.245
11.020	5.2	0.658	11.240
11.118	5.4	0.605	11.238
11.203	5.6	0.552	11.229
11.290	5.8	0.499	11.225
11.372	6.0	0.447	11.220
11.452	6.2	0.396	11.209
11.534	6.4	0.346	11.202
11.618	6.6	0.297	11.190
11.700	6.8	0.249	11.170
11.788	7.0	0.205	11.150
		pK_a(OH) =	11.21 (mean)
		Standard Deviation =	0.03

TABLE IV

2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE

(initial volume = 100 ml, 25.0°C, $[HClO_4]$ = 0.00986, [NaOH] = 0.1254, 0.983 g reagent, 50% (v/v) aqueous dioxane)

pH	VOLB	$\bar{p}$	pK_a
2.658	0.0	1.587	2.746
2.688	0.2	1.567	2.736
2.741	0.5	1.538	2.739
2.797	0.8	1.507	2.738
2.858	1.1	1.473	2.739
2.900	1.3	1.447	2.736
2.947	1.5	1.422	2.738
2.997	1.7	1.396	2.739
3.052	1.9	1.367	2.742
3.109	2.1	1.338	2.743
3.208	2.4	1.291	2.748
3.288	2.6	1.259	2.759
3.374	2.8	1.225	2.764

pK_a(NH) = 2.74 (mean)

Standard Deviation = 0.01

pH	VOLB	$\bar{p}$	pK_a
10.085	5.1	0.650	10.289
10.194	5.3	0.589	10.287
10.248	5.4	0.559	10.287
10.348	5.6	0.499	10.283
10.400	5.7	0.469	10.283
10.500	5.9	0.410	10.277
10.552	6.0	0.380	10.275
10.662	6.2	0.321	10.272
10.719	6.3	0.290	10.266
10.840	6.5	0.232	10.257
10.904	6.6	0.203	10.247

quinolinic pK_a(OH) = 10.28 (mean)

Standard Deviation = 0.01

TABLE V

2-HYDRAZINO-8-HYDROXYQUINOLINE

initial volume = 100 ml, 25.0°C, $[HClO_4]$ = 0.01044, [NaOH] = 0.1254, 0.0833 g reagent, 50% (v/v) aqueous dioxane)

pH	VOLB	$\bar{p}$	pK_a
6.053	1.2	1.781	6.540
6.198	1.4	1.730	6.561
6.305	1.6	1.680	6.559
6.411	1.8	1.623	6.564
6.566	2.1	1.543	6.576
6.655	2.3	1.491	6.574
6.808	2.6	1.412	6.589
6.912	2.8	1.360	6.596
7.087	3.1	1.279	6.611
7.214	3.3	1.227	6.618

hydrazino pK_a(NH) = 6.58 (mean)

Standard Deviation = 0.02

11.448	5.1	0.777	11.932
11.592	5.4	0.708	11.922
11.669	5.6	0.663	11.907
11.777	5.9	0.595	11.890
11.839	6.1	0.551	11.874
11.938	6.4	0.488	11.868
11.998	6.6	0.447	11.858
12.088	6.9	0.389	11.848
12.170	7.2	0.333	11.829
12.270	7.6	0.265	11.800
12.343	7.9	0.218	11.771

pK_a(OH) = 11.86 (mean)

Standard Deviation = 0.05

Representative Data for Formation Constants

TABLE VI

2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE AND NICKEL(II)

(initial volume = 110 ml, 25.0°C, [$HClO_4$] = 0.01005, [NaOH] = 0.1102, [Ni^{2+}] = 0.01152, 50% (v/v) aqueous dioxane)

(a) 0.05242 g ligand, 5 ml Ni^{2+} solution

pH	VOLB	pH	VOLB
2.493	0.0	2.528	0.3
2.566	0.6	2.602	0.9
2.648	1.2	2.694	1.5
2.750	1.8	2.811	2.1
2.855	2.3	2.903	2.5
2.958	2.7	3.020	2.9
3.084	3.1	3.158	3.3
3.248	3.5	3.294	3.6
3.400	3.8	3.463	3.9
3.530	4.0	3.601	4.1
3.678	4.2	3.759	4.3
3.758	4.4	3.957	4.5
4.080	4.6	4.225	4.7
4.379	4.8	4.642	4.9
5.052	5.0	5.652	5.1
6.032	5.2	6.312	5.3
6.568	5.4	6.918	5.5

(b) 0.02730 g ligand, 5 ml Ni^{2+} solution

2.419	0.0	2.440	0.3
2.472	0.6	2.508	0.9
2.544	1.2	2.582	1.5
2.623	1.8	2.671	2.1
2.720	2.4	2.758	2.6
2.799	2.8	2.840	3.0
2.887	3.2	2.937	3.4
2.997	3.6	3.033	3.7
3.071	3.8	3.111	3.9
3.151	4.0	3.200	4.1
3.248	4.2	3.306	4.3
3.373	4.4	3.448	4.5
3.543	4.6	3.669	4.7
3.839	4.8	4.090	4.9
4.640	5.0	5.645	5.11
6.015	5.2	6.292	5.3
6.550	5.4	6.899	5.5

(cont'd.)

Table VI (cont'd)

(c) 0.02379 g ligand, 10 ml Ni^{2+} solution

pH	VOLB	pH	VOLB
2.454	0.0	2.487	0.4
2.531	0.8	2.567	1.1
2.602	1.4	2.643	1.7
2.689	2.0	2.737	2.3
2.792	2.6	2.853	2.9
2.899	3.1	2.950	3.3
3.008	3.5	3.077	3.7
3.109	3.8	3.149	3.9
3.193	4.0	3.241	4.1
3.294	4.2	3.352	4.3
3.423	4.4	3.505	4.5
3.605	4.6	3.729	4.7
3.903	4.8	4.147	4.9

Complex	log β	Standard Deviation
NiL^+	11.8	0.2
NiL_2	24.0	0.1
$NiHL^{2+}$	16.62	0.09
$NiH_2L_2^{2+}$	32.8	0.1
$NiHL_2$	30.40	0.09

Standard Deviation in Titre = 35×10^{-3}

TABLE VII

2-HYDRAZINO-8-HYDROXYQUINOLINE AND NICKEL (II)

(initial volume = 110 ml, 25.0°C, $[HClO_4]$ = 0.01005, [NaOH] = 0.1102 $[Ni^{2+}]$ = 0.01152, 50% (v/v) aqueous dioxane)

(a) 0.04910 g ligand, 5 ml Ni^{2+} solution

pH	VOLB	pH	VOLB
3.632	2.2	3.699	2.3
3.767	2.4	3.837	2.5
3.902	2.6	3.978	2.7
4.053	2.8	4.128	2.9
4.209	3.0	4.300	3.1
4.397	3.2	4.519	3.3
4.671	3.4	4.897	3.5
5.215	3.6	5.571	3.7
5.849	3.8	6.053	3.9
6.212	4.0	6.350	4.1
6.473	4.2	6.588	4.3
6.710	4.41	6.808	4.5
6.925	4.6	7.040	4.7
7.161	4.8	7.298	4.9
7.439	5.0	7.588	5.1

(b) 0.02137 g ligand, 5 ml Ni^{2+} solution

pH	VOLB	pH	VOLB
3.532	3.2	3.611	3.3
3.705	3.4	3.800	3.5
3.898	3.6	3.992	3.7
4.088	3.8	4.180	3.9
4.265	4.0	4.350	4.1
4.432	4.2	4.515	4.3
4.607	4.4	4.702	4.5
4.817	4.6	4.945	4.7
5.112	4.8	5.349	4.9
5.900	5.0	5.943	5.1
6.448	5.2	6.713	5.3

(cont'd)

(c) 0.02063 g ligand, 10 ml Ni^{2+} solution

pH	VOLB	pH	VOLB
3.638	3.42	3.705	3.51
3.778	3.6	3.863	3.7
3.950	3.8	4.031	3.9
4.118	4.0	4.200	4.1
4.280	4.2	4.362	4.3
4.468	4.42	4.539	4.5
4.640	4.6	4.760	4.7
4.904	4.8	5.100	4.9
5.427	5.0	6.130	5.1
6.949	5.2	7.339	5.3

Complex	log β	σ
NiL^+	11.3	0.2
NiL_2	22.72	0.06
$NiHL^{2+}$	16.9	0.2
$NiH_2L_2^{2+}$	34.79	0.04
$NiHL_2$	30.48	0.03

Standard Deviation in Titre = 36 x 10^{-3}

TABLE VIII

2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE AND NICKEL(II)

(initial volume = 110 ml, 25.0°C, [$HClO_4$] = 0.01027, [NaOH] = 0.1243, 0.04847 g ligand, 50% (v/v) aqueous dioxane, 5 ml 0.01152 M Ni^{2+})

pH	VOLB	$\bar{n}$	pHL	log K(ML)
5.397	4.30	0.191	7.731	1.760
5.489	4.36	0.254	7.647	1.761
5.568	4.42	0.317	7.577	1.747
5.639	4.48	0.379	7.515	1.732
5.704	4.54	0.445	7.460	1.730
5.772	4.60	0.510	7.401	1.715
5.838	4.66	0.574	7.344	1.705
5.908	4.72	0.638	7.284	1.692
6.002	4.80	0.724	7.204	1.690
6.091	4.86	0.788	7.125	1.674

log K(ML) = 1.72 (mean)

Standard Deviation = 0.03

TABLE IX

8-HYDROXYQUINOLINE AND NICKEL

(initial volume = 110 ml, 25.0°C, $[HClO_4]$ = 0.01044, [NaOH] = 0.1243, 0.04308 g ligand, 50% (v/v) aqueous dioxane, 5 ml 0.01152 M Ni^{2+})

pH	VOLB	$\bar{n}$	pL^-
3.529	2.1	0.227	11.006
3.576	2.2	0.264	10.924
3.619	2.3	0.311	10.851
3.663	2.4	0.357	10.778
3.704	2.5	0.409	10.710
3.749	2.6	0.455	10.637
3.788	2.7	0.512	10.575
3.823	2.8	0.577	10.522
3.862	2.9	0.637	10.462
3.899	3.0	0.701	10.407
3.938	3.1	0.762	10.349
3.976	3.2	0.827	10.293
4.013	3.3	0.895	10.241
4.054	3.4	0.957	10.182
4.095	3.5	1.021	10.125
4.148	3.6	1.067	10.048
4.183	3.7	1.147	10.004
4.232	3.8	1.205	9.938
4.288	3.9	1.255	9.862
4.346	4.0	1.306	9.784

logK(ML) = 10.44

logK(ML_2) = 9.79

Relative standard deviation of calculated $\bar{n}$ values from experimental $\bar{n}$ values = 3.5×10^{-2}.

Representative data for the titration of other metal ions are of the same form as the data above and have not been tabulated. The titration curves are given, however, in RESULTS AND DISCUSSION.

APPENDIX IV

ULTRAVIOLET-VISIBLE ABSORPTION SPECTRA OF 2-PyOx, 2-PhOx and 2-HyOx

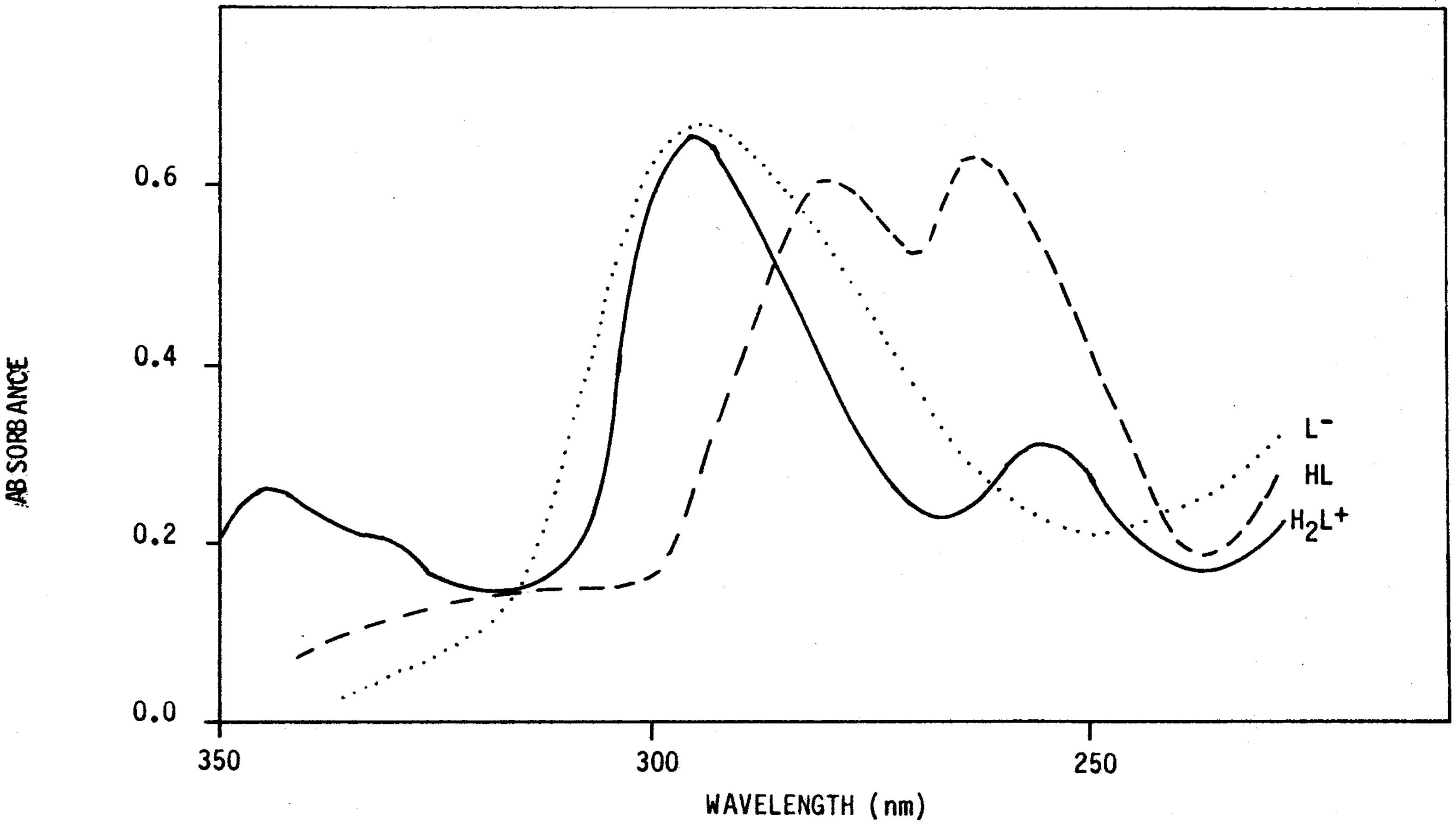


Figure 1. ABSORPTION SPECTRA OF ANIONIC (pH 13), CATIONIC (pH 2) AND NEUTRAL (pH 7) FORMS OF 2-(2'-PYRIDYL)-8-HYDROXYQUINOLINE.

Concentration = 2.06 x 10^{-5}M.

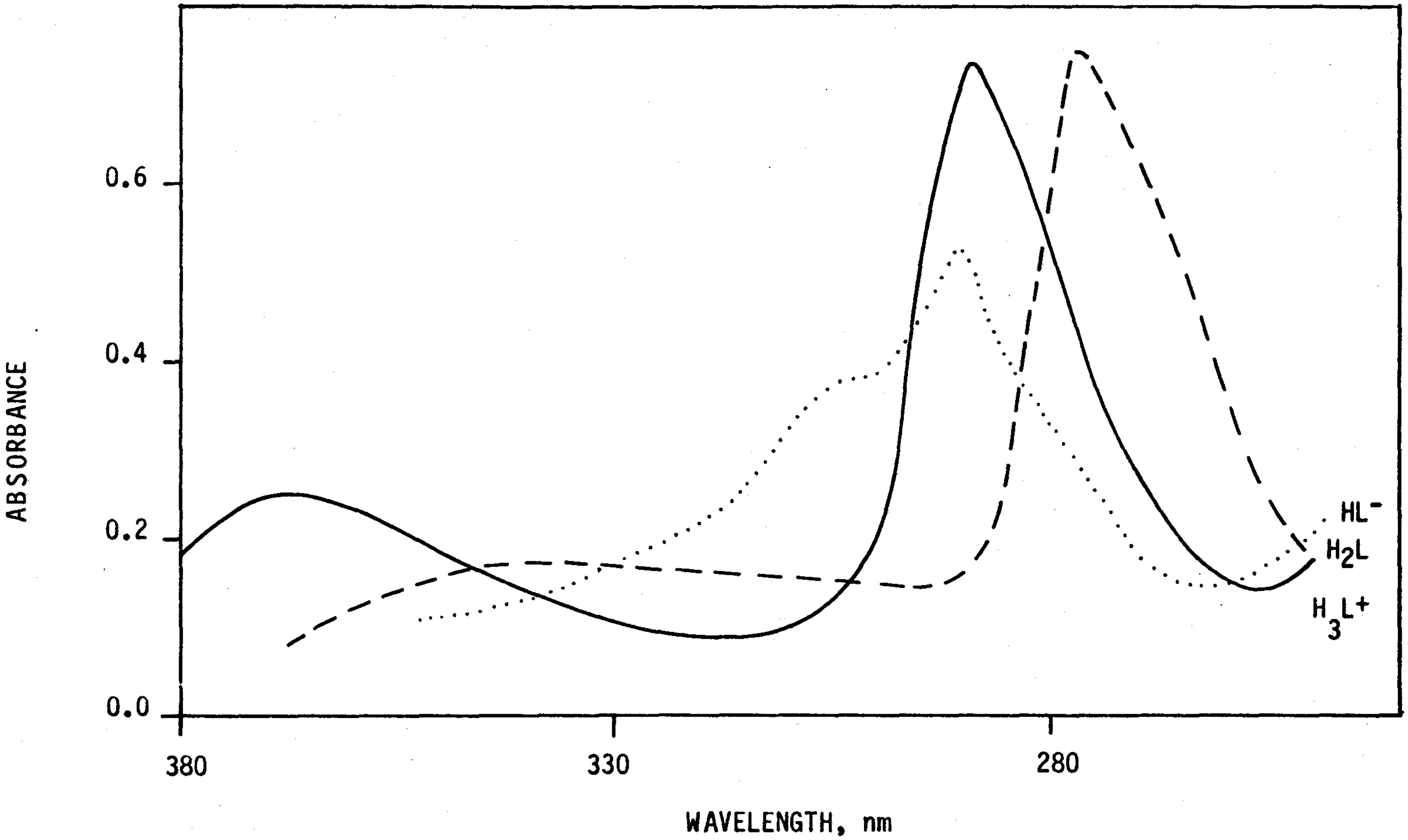


Figure 2. ABSORPTION SPECTRA OF ANIONIC (pH 12), CATIONIC (pH 1) AND NEUTRAL (pH 7) FORMS OF 2-(2'-HYDROXYPHENYL)-8-HYDROXYQUINOLINE.

Concentration = 1.77×10^{-5}M.

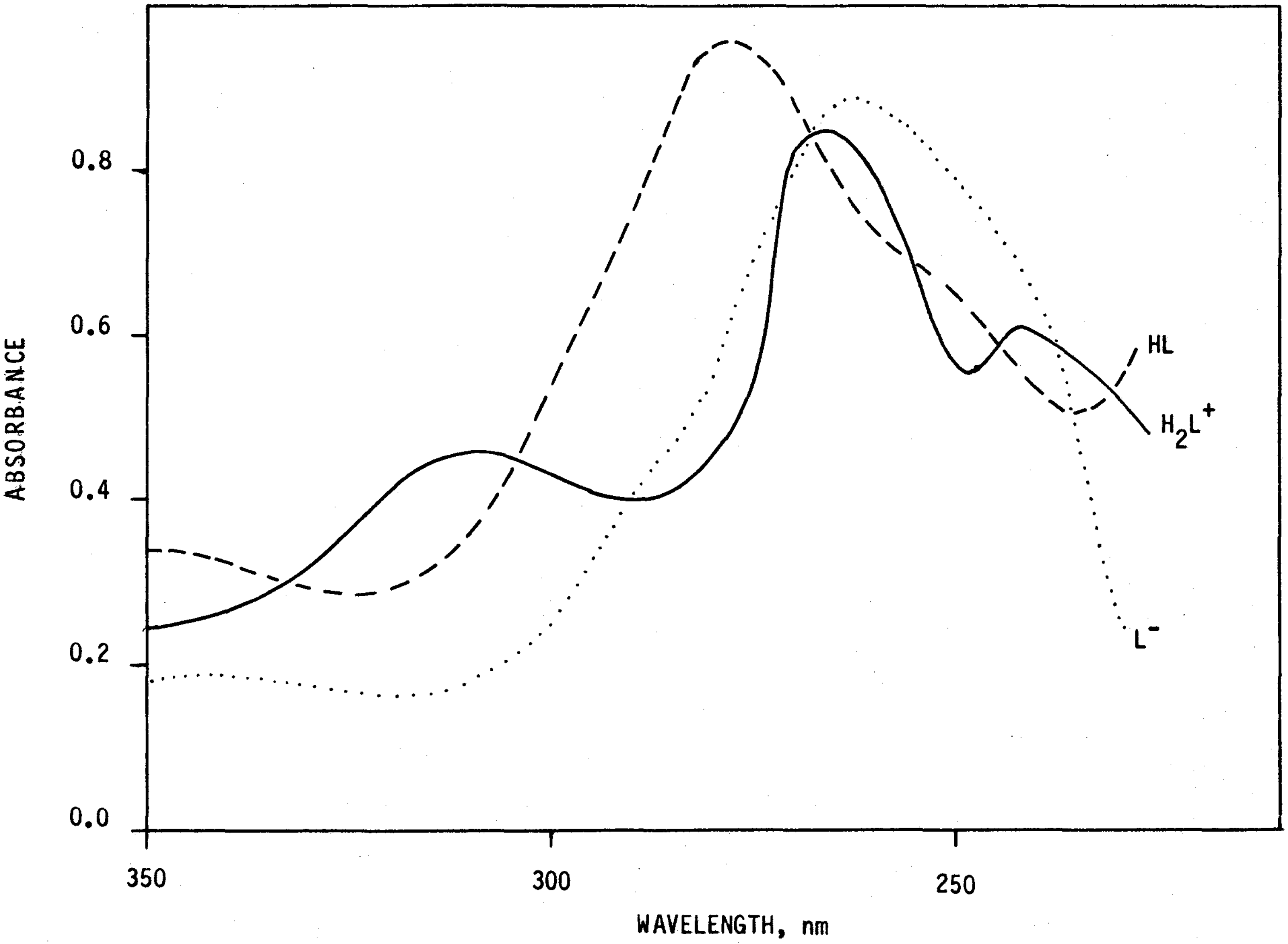


Figure 3. ABSORPTION SPECTRA OF ANIONIC (pH 12), CATIONIC (pH 3) AND NEUTRAL (pH 9) FORMS OF 2-HYDRAZINO-8-HYDROXYQUINOLINE.

Concentration = 2.03×10^{-5}M.

REFERENCES

1. A. E. Martell and M. Calvin, "Chemistry of the Metal Chelate Compounds", Prentice-Hall, Englewood Cliffs, N.J., 1956.

2. A. E. Martell and S. Chaberek, "Organic Sequestering Agents", John Wiley and Sons, New York, 1959.

3. F. P. Dwyer and D. P. Mellor, "Chelating Agents and Metal Chelates", Academic Press, New York, 1964.

4. F. Feigl, "Specific, Selective and Sensitive Reactions", Academic Press, New York, 1949, p.12.

5. F. Feigl, "Spot Tests", Elsevier, New York, 1958.

6. A. I. Vogel, "Quantitative Inorganic Analysis", Longmans, Green and Co., Toronto, 1951.

7. H. A. Flaschka, "EDTA Titrations", Pergamon Press, New York, 1964.

8. J. B. Headridge, "Photometric Titrations", Pergamon Press, New York, 1961.

9. O. Samuelson, "Ion Exchangers in Analytical Chemistry", John Wiley and Sons, New York, 1953.

10. G. H. Morrison and H. Freiser, "Solvent Extraction in Analytical Chemistry", John Wiley and Sons, New York, 1957.

11. F. J. Rossotti in J. Lewis and R. G. Wilkins, ed., "Modern Co-ordination Chemistry", Interscience, New York, 1960.

12. S. Ahrland, J. Chatt and N. R. Davies, Quart. Revs. (London) 12, 265 (1958).

13. R. J. Williams and J. D. Hale, Structure and Bonding, 1 , 249 (1966).

14. S. Ahrland, Structure and Bonding, 1 , 207 (1966).

15. R. G. Pearson, J. Am. Chem. Soc., 85, 3533 (1963).

16. R. S. Evans and J. E. Huhey, J. Inorg. Nucl. Chem., 32, 373; 383; 391 (1970).

17. H. Irving and H. Rossotti, Acta Chem. Scand., 10, 72 (1956).

18. R. A. Close and T. S. West, Talanta, 5, 221 (1960).

19. G.F. Smith and W. H. McCurdy, Anal. Chem., <u>24</u>, 371 (1952).

20. T. Tschugaeff, Z. anorg. Chem., <u>46</u> , 144 (1905).

21. R. G. Hollingshead, "Oxine and its Derivatives", Volumes I-IV, Buttersworth Scientific Publications, London, 1954.

22. L. Merritt and J. K. Walker, Ind. Eng. Chem. Anal. Ed., 16, 387 (1944).

23. J. P. Phillips and H. P. Price, J. Am. Chem. Soc., <u>73</u>, 4414 (1951).

24. H. Irving and D. L. Pettit in P. W. West, A. M. MacDonald and T. S. West, ed., "Analytical Chemistry 1962, Proceedings of the International Symposium, Birmingham University (U.K.) April, 1962", Elsevier, Amsterdam, 1963.

25. G. R. Horton, Anal. Chem., <u>39</u>, 1036 (1967).

26. P. L. Scherer and Q. Fernando, Chem. Comm., 1107 (1967).

27. P. L. Scherer and Q. Fernando, Anal. Chem., <u>40</u>, 1938 (1968).

28. Q. Fernando and Y. Kushi, Chem. Comm., 555 (1969).

29. T. J. Cardwell, Inorg. Nucl. Chem. Letters, <u>5</u>, 409 (1969).

30. W. D. Johnston and H. Freiser, Anal. Chim. Acta, <u>11</u>, 201 (1954).

31. H. Kaneko and K. Ueno, Bull. Chem. Soc. Japan, <u>39</u>, 1910 (1966).

32. A. Corsini and E. J. Billo, Can. J. Chem., <u>47</u>, 4655 (1969).

33. P. G. Manning, Can. J. Chem., <u>44</u>, 3057 (1966).

34. A. Corsini and E. J. Billo, J. Inorg. Nucl. Chem., <u>32</u>, 1241 (1970).

35. R. L. Stevenson and H. Freiser, Anal. Chem., <u>39</u>, 1354 (1967).

36. N. Bjerrum, Kgl. Danske Videnskab. Skrifter, Naturvidenskab. math. Afdel (1), 12 (1925).

37. J. C. Sullivan, J. Rydberg and W. F. Miller, Acta Chem. Scand., <u>13</u>, 2023 (1959).

38. H. Irving and H. S. Rossotti, J. Chem. Soc., 3397 (1953).

39. C. W. Childs, P. S. Hallman and D. D. Perrin, Talanta, <u>16</u>, 1119 (1969).

40. I. G. Sayce, Talanta, <u>15</u>, 1397 (1968).

41. M. Calvin and K. W. Wilson, J. Am. Chem. Soc., <u>67</u>, 2003 (1945).

42. L. G. Van Uitert and C. G. Haas, J. Am. Chem. Soc., 75, 451 (1953).

43. D. A. Skoog and D. M. West, "Fundamentals of Analytical Chemistry", Holt, Rhinehart and Winston, 1963, p. 347.

44. G. Schwarzenbach, "Complexometric Titrations", Interscience Pub., New York, 1957.

45. B. Karlberg and G. Johansson, Talanta, 16, 1545 (1969

46. H. Irving and H. S. Rossotti, J. Chem. Soc., 2910 (1954).

47. H. M. Irving and U. S. Mahnot, J. Inorg. Nucl. Chem., 30, 1215 (1968).

48. G. R. Nash and C. B. Monk, Trans. Faraday Soc., 54, 1657 (1959).

49. H. S. Harned and L. D. Fallon, J. Am. Chem. Soc., 61, 1374 (19 9).

50. E. A. Unwin, R. G. Beamer and Q. Fernando, Anal. Chem. Acta, 39, 95 (1967).

51. J. P. Phillips, R. L. Elbinger and L. L. Merritt, J. Am. Chem. Soc., 71, 3989 (1949).

52. F. Sorm and J. Sicher, Coll. Czech. Chem. Comm., 14, 331 (1949).

53. H. Gilman, W. Gregory and S. Spatz, I. Org. Chem., 16, 1788 (1951).

54. H. Gilman and C. Arntzen, J. Am. Chem. Soc., 69, 1537 (1947).

55. J. K. Towle, Iowa State College Journal of Science, 26, 308 (1951-1952).

56. T. A. Geissman, J. Org. Chem., 11, 741 (1946).

57. M. Kolobielski, J. Heterocyclic Chem., 3, 275 (1966).

58. T. Rudolph, F. Prystal and J. P. Phillips, J. Med. Pharm. Chem., 10, 981 (1967).

59. M.Thompson, Ph.D. Thesis, McMaster University, 1970.

60. R. F. Hirch and J. A. Sbanbaro, Analytical Letters, 1, 385 (1968).

61. B. C. Baker and D. T. Sawyer, Anal. Chem., 40, 1945 (1968).

62. H. T. Britton and W. G. Williams, J. Chem. Soc., 96 (1936).

63. R. Näsänen, P. Lumme and A. L. Mukula, Acta Chem. Scand., 5, 1199 (1952).

64. I. R. Beattie and M. Webster, J. Phys. Chem., 66, 115 (1962).

65. G. Anderegg, Helv. Chim. Acta, 46, 2397 (1963).

66. W. D. Johnston and H. Freiser, J. Am. Chem. Soc., 74, 5239 (1952).

67. Norio Yui, Bull. Inst. Phys. Chem. Research (Tokyo), 20, 256 (1941); C.A., 35, 4660 (1941).

68. M. A. Khwaja, T. J. Cardwell and R. J. Magee, Anal. Chim. Acta, 53, 317 (1971).

69. M. Thompson, J. Abraham and A. Corsini, Talanta, 18, 481 (1971).

70. A. Corsini and E. J. Billo, J. Inorg. Nucl. Chem., 32, 1249 (1970).

71. R. C. Turner, Can. J. Chem., 47, 2521 (1969).

72. C. Brosset, Acta Chem. Scand., 6, 910 (1952).

73. C. Brosset, G. Biedermann and L. G. Sillen, ibid, 8, 1917 (1954).

74. L. G. Sillen, Quart. Rev., 13, 146 (1959).

75. E. Matijevic and B. Tezak, J. Phys. Chem., 57, 951 (1953).

76. E. Matijevic, K. G. Mathoi, R. H. Ottewitt and M. Kerker, ibid, 65, 826 (1961).

77. J. Faucherre, Bull. Soc. chim. France, 21, 253 (1954).

78. H. Kubota, Diss. Abs., 16, 864 (1956).

79. J. Aveston, J. Chem. Soc., A, 4438 (1965).

80. J. Bjerrum, "Metal Ammine Formation In Aqueous Solution", P. Haase and Son, Copenhagen, 1941.

81. A. W. Adamson, J. Am. Chem. Soc., 76, 1578 (1954).

82. J. Agterdenbos, J. Chem. Ed., 45, 230 (1968).

83. A. K. Babko, Talanta, 15, 721 (1968).

84. A. Corsini, Q. Fernando and H. Freiser, Inorg. Chem., 2, 224 (1963).

85. "Stability Constants", Special Publication No.17, London Chemical Society, 783, 1964.

86. R. C. Voter, C. V. Banks, V. A. Fassel and P. W. Kehres, Anal. Chem., 23, 1730 (1951).

87. R. E. Rundle and M. Parasol, J. Chem. Phys., 20, 1487 (1952).

88. R. E. Rundle, Acta Cryst., 6, 487 (1953).

89. R. E. Rundle, J. Phys. Chem., 61, 45 (1957).

90. R. Blinc and D. Hadzi, J. Chem. Soc., 4536 (1958).

91. R. G. Charles and H. Freiser, Anal. Chim. Acta, 11, 101 (1954).

92. L. G. Van Uitert and W. C. Fernelius, J. Am. Chem. Soc., 76, 375 (1954).

93. K. Burger and I. Ruff, Talanta, 10, 329 (1963).

94. C. Banks and S. Anderson, Inorg. Chem., 2, 112 (1963).

95. Y. Yamano, I. Masuda and K. Shinra, Inorg. Nucl. Chem. Letters, 4, 581 (1964).

96. K. Burger, I. Ruff and F. Ruff, Proceedings of the Symposium on Coordination Chemistry, Hungary, 1964, Edited by M. T. Beck, Budapest, Akadémiai Kiado, 1965.

97. B. Egneus, Anal. Chim. Acta, 48, 291 (1969).

98. E. Cartmell and G. Fowles, "Valency and Molecular Structure", Butterworths, London, 1966, p. 140.

99. T. R. Harkins and H. Freiser, J. Am. Chem. Soc.,78, 1143 (1956).

100, S. P. Bag, Q. Fernando and H. Freiser, Inorg. Chem., 3, 93 (1969).

101. R. W. Green, P. S. Hallman and F. Lions, Inorg. Chem., 3, 376 (1964).

102. G. Beech, Quart. Rev., 23, 410 (1969).

103. E. R. Clarke and A. E. Martell, J. Inorg. Nucl. Chem., 32, 911 (1970).

104. B. M. Dahl and O. Dahl, Acta Chem. Scand., 23, 1503 (1969).

105. V. N. Kumok and V. V. Serebrennikov, Russ. J. Inorg. Chem., 11, 1488 (1966).

106. W. A. Johnson and R. G. Wilkins, Inorg. Chem., 9, 1917 (1970)

107. Y. Wormser, Bull. Soc. Chim. Fr., 21, 387 (1954).

108. C. J. Ballhausen, "Introduction to Ligand Field Theory", McGraw-Hill, Toronto, 1962, p. 268-273.

109. A. D. Liehr, J. Phys. Chem., 64, 43 (1960).

110. H. Irving and D. H. Mellor, J. Chem. Soc., A, 5222 (1962).

111. P. L. Pecsok and J. Bjerrum, Acta Chem. Scand., 11, 1419 (1957).

112. B. R. James and R. J. Williams, J. Chem. Soc., A, 2007 (1961).

113. C. K. Jørgensen, Acta Chem. Scand., 9, 1362 (1965).

114. F. J. C. Rossotti and H. S. Rossotti, "The Determination of Stability Constants", McGraw-Hill, New York, 1961.

115. H. Steger, "Solution Stability Relationships of Metal 8-Hydroxyquinolinates", Ph.D. Thesis, McMaster University, August 1969.

116. F. E. King and J. A. Sherred, J. Chem. Soc., 415 (1942).

117. H. Gilman, J. A. Beel, M. W. Bullock, G. E. Dunn and L. S. Miller, J. Am. Chem. Soc., 71, 1499 (1949).

118. H. Gilman and A. H. Haubein, J. Am. Chem. Soc., 66, 1515 (1944).

119. H. Gilman and S. M. Spatz, J. Org. Chem., 16, 1485 (1951).

120. E. Klingsberd, "Heterocyclic Compounds; Pyridine and its Derivatives", Interscience, New York, 1961, Part 2, p. 422.

121. J. P. Phillips, E. M. Barrall and R. Breese, Trans. Kentucky Acad. Sci., 17, 135 (1956); C. A., 51, 11349a (1958).

122. M. Hamana and K. Funakoshi, Yakugaku Zasshi, 84, 24 (1964); C. A., 61, 3068 (1964).

123. I. G. Sayce, personal communication.