[CONTRIBUTION NO. 836 FROM THE DEPARTMENT OF CHEMISTRY, UNIVERSITY OF PITTSBURGH]

Structure and Behavior of Organic Analytical Reagents. I. The Calvin–Bjerrum Method for Determination of Chelate Stability

BY HENRY FREISER, ROBERT G. CHARLES AND W. DWIGHT JOHNSTON

The Calvin–Bjerrum technique for evaluation of chelate stability has been adapted to the study of chelating agents possessing more than one acidic or basic group. The general applicability of the method is reviewed. In this connection, hydrolysis of some metal ions was studied as a possible source of error in the method.

Introduction

Great advances have been made in recent years toward the elucidation of the relation between the structure of organic chelating reagents and their analytical behavior. While the aim of developing reagents of such selectivity that they would react with only one metal each is still beyond reach, much has been learned to help design better organic reagents for use in difficult metal separations.

One of the most useful ways to characterize an organic chelating reagent is by the determination of the stabilities of its metallic chelates. The first quantitative evaluation of chelate stability was performed by Calvin and Wilson[1] in 1945 by an adaptation of the method used by Bjerrum[2] in 1941 to study the formation constants of metal ammines. This method, applicable to reagents having acidic or basic properties, involves following the course of chelate formation by measuring the amount of hydrogen ion released in the reaction

$$M^{2+} + 2HR \rightleftharpoons MR_2 + 2H^+$$

where M refers to a metal ion, HR refers to the chelating agent and MR_2 refers to the metallic chelate.

At the University of Pittsburgh, there has been initiated an extensive program to attack the many vital problems concerning the relation of stability and structure in organic analytical reagents. Some of the questions to be considered include, (a) the extent to which steric factors such as ring size or the presence of blocking groups in reagents affect stability of their chelates; (b) the effect upon stability of replacing oxygen atoms in the reactive grouping of a reagent by sulfur or nitrogen atoms; (c) determining the nature of the structural parameters of metals which affect the order of the stability of their chelates; and (d) the search for reagents in which the order of the metals mentioned in (c) is unusual, *i.e.*, for "specific" reagents.

Several other methods have been employed to obtain information concerning relative stabilities of chelates. Shifts of the polarographic half-wave potentials of related chelate compounds have been taken as indications of their relative stabilities by Calvin and Bailes[3] who base their conclusions on the assumption that the more stable a metallic chelate the more difficult it would be to reduce. In another instance,[4] the rate of exchange of radioactive metal ions with the same metal in chelate combination was assumed to be related to chelate stability. Chelates which experience almost instantaneous exchange were thought to contain an ionically bound metal, while a slow rate of metal exchange was considered an indication of the covalent character of the organo-metallic bond and hence, of its greater strength. Merritt and Phillips[5] employed still another method, consisting of noting the value of the *p*H of precipitation of similar metallic chelates. The *p*H of precipitation is related to chelate stability, but as it also involves solubility phenomena, it cannot be considered as an unequivocal basis for stability evaluation.

The Calvin–Bjerrum titration method appears to be the most reliable for chelate stability evaluation inasmuch as it appears to be the only one of those mentioned which permits quantitative determination of formation constants. In this paper, the method is adapted to the study of reagents possessing more than one acidic or basic group. The phenomenon of hydrolysis of metals, under conditions paralleling those encountered in the operation of the method has been investigated and its bearing on the general applicability of the method considered.

Experimental

Materials.—Carbide and Carbon Co. dioxane was purified by refluxing with metallic sodium for 12 hours and then fractionated through a four foot, glass helices-packed column. The material used in this study boiled over a range of 0.1°.

A time-temperature cooling curve of this material revealed that the temperature range from the time the first crystals appeared to almost complete solidification, was less than 0.1°. The material had a melting point of 11.63°.

Stock solutions of metal ions of approximately 0.01 molarity were prepared by dissolving the corresponding reagent grade perchlorates (G. Frederick Smith Chemical Co.) in water. The copper(II) and cobalt(II) solutions were analyzed by electrodeposition. The nickel(II) solution was standardized by precipitation with dimethylglyoxime. The lanthanum(III), magnesium(II), cadmium(II), and aluminum(III) solutions were determined by precipitation with 8-hydroxyquinoline. The cerium(III) solution was analysed by precipitation with 8-hydroxyquinoline and subsequent ignition to CeO_2. The lead(II), zinc(II), and manganese(II) solutions were standardized gravimetrically as $PbSO_4$, $ZnNH_4PO_4$, and $Mn_2P_2O_7$, respectively.

The approximately 0.1 *N* sodium hydroxide was carbonate-free and standardized with Mallinkrodt reagent grade potassium acid phthalate. Perchloric acid of about 0.01 normality was prepared by dilution of Fisher 70% reagent acid and standardized by comparison with the sodium hydroxide solution.

Apparatus.—The titration apparatus consisted essentially of a water-jacketed Pyrex vessel loosely fitted with a clear Bakelite cover, through which could be inserted the tips of two burets, a thermometer, a nitrogen inlet tube and the extension electrodes of a Model G Beckman *p*H meter. By circulating water from a constant temperature bath through the water jacket, the temperature of the solution to be measured was kept at 25 ± 0.3°. A carbon dioxide-free nitrogen

---

[1] M. Calvin and K. W. Wilson, THIS JOURNAL, **67**, 2003 (1945).

[2] J. Bjerrum, "Metal Ammine Formation in Aqueous Solution," P. Haase & Son, Copenhagen, 1941.

[3] M. Calvin and R. H. Bailes, THIS JOURNAL, **68**, 949 (1946).

[4] R. B. Duffield and M. Calvin, *ibid.*, **68**, 557 (1946).

[5] L. L. Merritt and J. P. Phillips, *Record Chem. Progress* (Kresge-Hooker Sci. Lab.), **10**, 59 (1949).

atmosphere was maintained above the solution. Rapid mixing was achieved by means of a magnetic stirrer which was stopped during the actual *p*H measurements. The *p*H meter was standardized with Beckman standard buffer solution (*p*H 7.00 ± 0.02) and with a buffer solution (*p*H 9.00 ± 0.02) prepared from a Coleman buffer tablet.

Titration Procedure.—Fifty milliliters of 0.01 *N* perchloric acid was pipetted into the titration vessel. To this was added by pipet a 50-ml. aliquot of a dioxane solution containing a weighed quantity of chelating agent, 5 ml. of the metal perchlorate solution, and 5 ml. of dioxane. As some of the reagents studied slowly decomposed in solution, the dioxane solution of the reagent was prepared as needed. The titration vessel and its contents were purged with nitrogen for five to ten minutes and the titration begun. The base was added in small increments which were further decreased in regions of rapidly changing *p*H values. With each addition of base, there was added an equal volume of dioxane in order to keep the dioxane content, and consequently, the dielectric constant, of the solvent constant. An atmosphere of nitrogen was maintained over the solution throughout the titrations.

Titrations of the metals in the absence of reagents were carried out in order to learn how seriously hydrolysis of the metallic cations might interfere with stability measurements. The results of these titrations are summarized graphically in Fig. 1. Despite the use of a nitrogen atmosphere true hydrolysis curves could not be obtained in the cases of manganese(II) and cerium(III) due to oxidation.

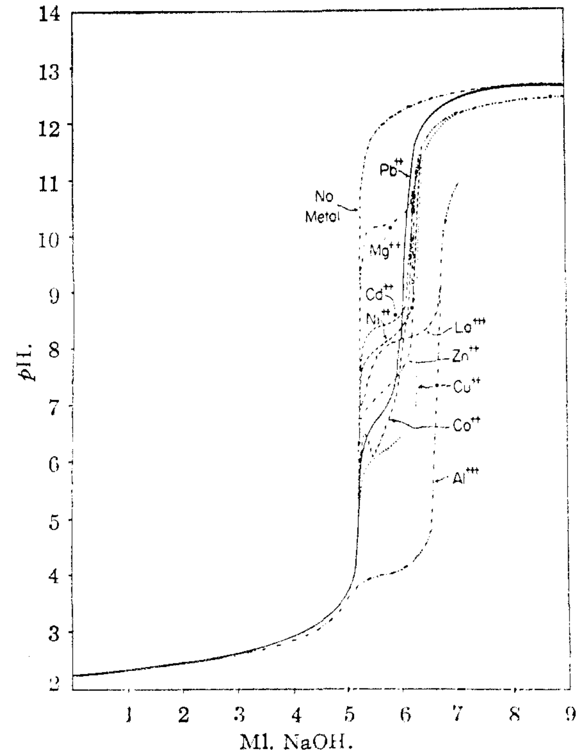

Fig. 1.—Hydrolysis curves in 50% water-dioxane at 25°.

Titrations of the reagents in the absence of metals served as the basis for the determination of their ionization constants.

In all calculations, the concentrations were corrected for the change in volume produced during the titrations.

Calculations of Chelate Stability Constants.—The calculations of stability constants for reagents such as 8-hydroxyquinoline which have in addition to a weakly acid group, a basic group can be carried out by a modification of the method used by Calvin and Wilson.[1] In the equations developed here, the acid dissociation constants of the protonated nitrogen K_{NH} and of the phenol, K_{OH}, are used. From the following equations expressing material balances of metal and reagent and charge balance

$$\text{Total metal concn., } T_M = M^{2+} + MR^+ + MR_2$$

$$\text{Total reag. concn., } T_{HR} = H_2R^+ + HR + R^- + MR^+ + 2MR_2$$

$$2M^{2+} + MR^+ + H^+ + Na^+ = ClO_4^- + R^- + OH^-$$

$$ClO_4^- = A + 2T_M$$

it is possible to obtain *n*, the average number of R^- bound to a M^{2+}.

$$n = \frac{1}{T_M}\left[T_{HR} - S\left(\frac{K_{NH} + H^+}{K_{NH} + 2H^+}\right)\right]$$

$$R^- = \frac{S \times K_{NH} \times K_{OH}}{H^+(K_{NH} + 2H^+)}$$

$$\text{where } S = T_{HR} + A - Na^+ + OH^- - H^+$$

These calculations can also be shown to apply without change to the case of trivalent metal ions. For reagents having two acid groups, one weak and one strong, and one basic group, such as 8-hydroxyquinoline sulfonic acid, the same equations can be shown to apply upon a redefinition of S as $S = 2T_{HR} + A - Na^+ + OH^- - H^+$.

The stepwise chelate formation constants K_1 and K_2 are obtained from the plot of $\bar{n}$ against p_R, at values of $\bar{n} = {}^1/_2$ and ${}^3/_2$, respectively.

Discussion

The titration curves obtained in the study of the hydrolysis of the metal salts were almost identical with those obtained by Britton,[6] indicating that the 50% dioxane solvent was not significantly different from water in this regard. In all cases except that of lead, two equivalents of base reacted with each metal ion. This signified that the hydroxide, not a basic salt, was the solid which precipitated. With lead, one and a half equivalents of base were consumed, up to a *p*H of 8 as observed in water,[6] indicating the formation of basic salt of the probable formula, $Pb(ClO_4)_2 \cdot 3Pb(OH)_2$. It is interesting to note in passing that, in 50% dioxane, acid–base titration curves undergo greater *p*H change in the end-point region than is the case in analogous aqueous titrations. The increased height of the vertical *p*H region, due probably to the decrease of the ion product of water in this medium, should be of advantage in titrating weaker acids than can be practicably titrated in water.

The *p*H values at which the metals begin to hydrolyze set an approximate upper limit to the *p*H range in which chelate stability measurements can be successfully carried out. Hydrolysis offers a serious interference since the release of hydrogen ions through this process would lead to the calculation of stability constants that would be too high. As can be seen from the formulation of the formation equilibrium as

$$M^{2+} + 2HR \rightleftharpoons MR_2 + 2H^+$$

$$K = \frac{(MR_2)(H^+)^2}{(M^{2+})(HR)^2}$$

it can be seen that the *p*H range in which chelate formation occurs can be lowered, and the extent to which hydrolysis interferes reduced, by increasing the concentration of HR, the chelating agent.

With the concentration of reagents to be used in this study, the amount of hydrogen ion released by chelate formation becomes small compared with the total amount of hydrogen at *p*H values below 3. Hence, the optimum *p*H range for accurate stability measurements is from about 3 to 6 (or somewhat higher, with metals like nickel which have less tendency to hydrolyze).

While concentrations rather than activities

---

(6) H. S. Britton, *J. Chem. Soc.*, **127**, 2110 (1925).

are employed in the calculations, the solutions are sufficiently dilute to make this error of minor significance. It is estimated that stability constant values determined as herein described are accurate within 0.2 log K unit.

Acknowledgment.—We wish to acknowledge gratefully the support of the Atomic Energy Commission for this work.

PITTSBURGH 13, PENNSYLVANIA

RECEIVED AUGUST 21, 1951

[CONTRIBUTION NO. 838 FROM THE DEPARTMENT OF CHEMISTRY, UNIVERSITY OF PITTSBURGH]

Structure and Behavior of Organic Analytical Reagents. II. Stability of Chelates of *o*-Aminophenol and of *o*-Aminobenzenethiol[1]

BY ROBERT G. CHARLES AND HENRY FREISER

The acid dissociation constants of *o*-aminophenol and *o*-aminobenzenethiol and the stability constants of the chelates of these reagents with some common divalent metals have been determined by potentiometric titration techniques. The chelates of *o*-aminophenol are less stable than those of *o*-aminobenzenethiol, despite the greater acidity of the latter reagent. This is attributed to the probably greater covalent character of the metal–sulfur bond than that of metal–oxygen bond. While the order of stability of metal chelates in *o*-aminophenol and *o*-aminobenzenethiol is similar to that observed in beta-diketones and salicylaldehydes, there are some significant differences. The reactions of the *o*-aminophenol and *o*-aminobenzenethiol with metals have been observed and tabulated.

Quite a few organic analytical reagents such as the α-amino acids, 8-hydroxyquinoline and quinaldic acid possess the reactive grouping, N–C–C–OH, capable of forming five-membered ringed chelate compounds with metals. For this reason, work on *o*-aminophenol as a model compound was undertaken. It was also desired to investigate the stabilities of Schiff bases of *o*-aminophenol and various aldehydes as compounds more closely analogous to 8-hydroxyquinoline, but unfortunately these compounds reverted almost immediately in solution to the original aminophenol and aldehyde. A comparison of the stabilities of the aminophenol complexes with those of aminobenzenethiol was undertaken in order to evaluate the effect of the substitution of sulfur for the more electronegative oxygen atom.

Experimental

Materials.—The purification of the dioxane and the preparation and standardization of metal perchlorate solutions and of standard sodium hydroxide and perchloric acid has been previously described.[1]

Eastman Kodak Co. practical grade *o*-aminophenol was decolorized with activated charcoal and recrystallized several times from water. It was found necessary to maintain an atmosphere of nitrogen over the hot phenol solution to prevent its oxidation, m.p. 175–176°. *o*-Aminobenzenethiol, provided in generous quantity by the American Cyanamid Co., was purified by fractional distillation under reduced pressure, b.p. 87.5–91.0°, at *ca* 1 mm. Since this compound is very sensitive to atmospheric oxidation, it was distilled shortly before use and kept in a frozen condition under a nitrogen atmosphere until used. Solutions of most of the common metallic ions containing 10 mg. metal per ml. were prepared from reagent grade salts. These were used only in the qualitative tests with the phenol and thiol.

Apparatus.—The apparatus is that previously described.[1]

Titration Procedure.—While the titration procedure used here has been described,[1] a modification employed in the aminobenzenethiol runs is noteworthy. In order to minimize handling and exposure to atmospheric oxygen, the approximate amount of the compound was introduced directly into the titration vessel without weighing. This amount was determined from the titration curve.

Results.—The results of representative titrations of metals with the two compounds are given in Table I along with their acid dissociation constants. It was found that chelate formation with aminophenol in ten fold ratio to metal ion occurred in the same pH region in which metal hydrolysis took place[1] (with the exception of copper). Since no precipitation was observed during these titrations, calculations of formation constants were carried out. Values of constants obtained from these calculations were different from those obtained from titrations in which the aminophenol was present in 100-fold excess in order to lower the pH range of chelate formation. A comparison of these values, presented in Table II, serves to illustrate the danger of accepting as valid, titration data obtained when pH of chelate formation is in the range in which metal ions hydrolyze appreciably.

The chelate formation constants of copper, nickel and cobalt *o*-aminobenzenethiolates could not be obtained from the titration data; because of their insolubility in the solvent these chelates precipitated early in the titration (see Fig. 1). As considerable fractions of the titrations in the case of zinc and lead proceeded in the absence of precipitation, it was

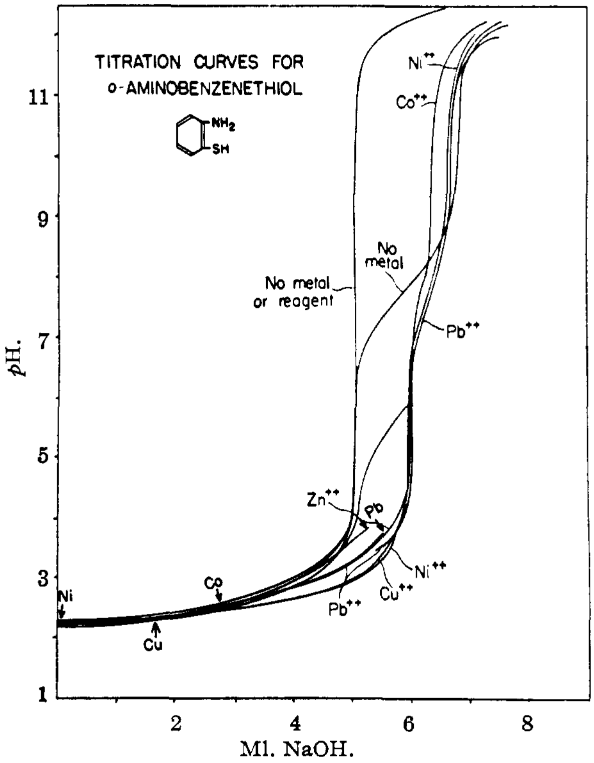

Fig. 1.—Arrow at point of appearance of ppt.

---

(1) For paper I see THIS JOURNAL, **74**, 1383 (1952).