SALICYLHYDROXAMIC ACID AND ITS IRON(III) COMPLEXES

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The dissociation constants of salicylhydroxamic acid (H_2L), $pK_{a1} = 6{\cdot}73$ and $pK_{a2} = 9{\cdot}15$ (20°C, $I = 1{\cdot}0$ ($NaClO_4$)), and the stability constant of the $FeHL^{2+}$ complex, $\log \beta = 9{\cdot}09$, were determined by the potentiometric and spectrophotometric methods. The low-soluble $Fe(OH)(HL)_2.3\,H_2O$ complex was also isolated. The reagent can be conveniently used as an indicator for the chelometric determination of iron.

Hydroxamic acids are important complexans which have found wide application in the analytical practice[1]. Previously we were concerned with the preparation of derivatives of aminopolycarboxylic acids containing the hydroxamic functional group and with the investigation of their chelating properties[2-5]. We also synthesized reagents that, in addition to the hydroxamic functional group, contained the azomethyne and phenolic functional groups[6,7], and we paid special attention to the application of the reagents as indicators for the chelometric determination of iron[7-9].

Salicylhydroxamic acid, whose synthesis has been first reported by Jeanreanaud[10], has been used as an analytical reagent for the photometric determination of vanadium, uranium, molybdenum and iron[11-13] as well as titanium[14] and some other metals[15]. The present study deals with the acid-base and complexing properties of this substance, with a view to employing this reagent as a chelometric indicator for iron.

EXPERIMENTAL

Apparatus and Chemicals

Spectrophotometric measurements were performed on a Unicam SP 8-200 instrument in 100 ml fused silica cells 3·48 cm optical pathlength; some measurements were carried out on a Unicam SP 500 manually operated spectrophotometer. Simultaneously with the absorbances, the pH

of the solutions was measured using a PHM-64 pH-meter equipped with a GK 2 401 B combined electrode (Radiometer, Copenhagen). Standard buffers served for calibration at pH 1·68, 4·00, 7·00, and 9·22. The pH adjustment was performed with 0·1M-NaOH or $HClO_4$ using an ABU-12 microburette (Radiometer, Copenhagen). The ionic strength was invariably adjusted to $I = 1·0$ with $NaClO_4$.

The analytical concentration of hydrogen ions was calculated, for solutions at pH < 1·5, from the volume and concentration of the perchloric acid added; for the remaining solutions, from the pH measured, using the experimental value of $\gamma_{H^+} = 0·840$ at $I = 1·0$ ($NaClO_4$).

Elemental analyses were carried out on a Carlo Erba 1102 instrument, infrared spectra in KBr disks were scanned on a Perkin–Elmer 377 spectrophotometer. Standard solution of $Fe(ClO_4)_3$, $c = 0·5$ mol l^{-1}, was prepared from the chemical of reagent grade purity (Fluka, Buchs); its concentration was determined chelometrically. Solution of salicylhydroxamic acid was prepared by dissolving the desired weighed amount in water. All the other chemicals used were of reagent grade purity.

Synthesis of Salicylhydroxamic Acid

Salicylhydroxamic acid was prepared by the modified method of Hurd and Botteron[16], who reported this procedure for arylhydroxamic acids. Methyl salicylate was reacted with hydroxylammonium chloride in aqueous–alcoholic solution in the presence of NaOH to obtain sodium salicylhydroxamate. The acid was obtained by acidification (pH 5—6) with intense cooling. The product was recrystallized from hot water; yield 62%, m.p. 169—170°C is in agreement with published data[10]. For $C_7H_7NO_3$ (152·2) calculated: 54·70% C, 4·56% H, 9·13% N; found: 54·65% C, 4·60% H, 9·10% N.

Isolation of Fe(III)-Salicylhydroxamic Acid Complex

The complex, which is low soluble in water and in ether, was synthesized following the procedure[17]: 3·36 g (22 mmol) of salicylhydroxamic acid (H_2L) is dissolved in 41·1 ml of 0·5M-NaOH, 17·65 ml of 0·5M-$Fe(ClO_4)_3$ is added, and the solution is adjusted to pH 7·5 with 0·5M-NaOH. The stirred mixture is heated at 50°C and cooled for 2 h. The precipitate separated is filtered out, washed with a small amount of water and ether, and dried in air and then in a dessicator over P_2O_5.

Iron was determined chelometrically after mineralization with HNO_3 and $HClO_4$. For $Fe(OH)(HL)_2.3\,H_2O$, $C_{14}H_{19}N_2O_{10}Fe$ (431·1), calculated: 39·00% C, 4·44% H, 6·49% N, 12·95% Fe; found: 38·90% C, 4·69% H, 6·42% N, 12·71% Fe. Water content was determined thermogravimetrically, calculated: 12·53% H_2O found: 12·24% H_2O.

Determination of Dissociation Constants of Salicylhydroxamic Acid

The dissociation constants of salicylhydroxamic acid were determined spectrophotometrically by simultaneous pH and absorbance (320 and 325 nm) measurements[5] for solutions at $c_L = 80$ μmol l^{-1} with gradual adjustment of pH with 0·1M-NaOH and constant nitrogen flushing. Alternatively, the constants were determined by potentiometric titration of the acid ($c_L = 4$ mmol. l^{-1}) with NaOH ($c = 0·1$ mol l^{-1}) at 20°C, applying nitrogen flushing.

The calculation of the constants from the various buffer regions of the potentiometric neutralization curve was performed using our program DISCO-2-POT for the TI-59 based on multiple linear regression[18]; from the spectrophotometric data the constants were derived by means of our programs GRAFAN and LOGAN for the TI-59.

Spectrophotometric Study of Fe(III)–Salicylhydroxamic Acid System

The complex formation in strongly acidic solutions was examined by measuring the absorption curves at pH = 0—2 in the region of 400—700 nm, using a five-fold concentration excess of iron ($c_L = 0{\cdot}25$ mmol l^{-1}, $c_{Fe} = 1{\cdot}25$ mmol l^{-1}) or of reagent ($c_L = 1{\cdot}25$ mmol l^{-1}, $c_{Fe} = 0{\cdot}25$ mmol l^{-1}). At pH 2—8·5 the solution becomes cloudy even if the acid is present in an excess, and a red-brown precipitate of the complex separates. The results were treated to obtain the molar absorptivities of $FeHL^{2+}$ and the complex formation constant. The complex composition was studied by using the Job method of isomolar solutions.

Chelometric Determination of Iron(III) Using Salicylhydroxamic Acid as Indicator

To 10 ml of solution containing 558·5 μg of Fe(III) in a titration flask was added 40 ml of water and 1 ml of 1% solution of salicylhydroxamic acid in a water–alcohol 1 : 1 mixture, and the system was adjusted to pH 1·5 with 1M-HNO_3 and titrated with Chelaton 3 from the violet to light-yellow colour.

RESULTS AND DISCUSSION

The identity of the salicylhydroxamic acid prepared was confirmed by elemental analysis and IR spectra; in the latter, attention was paid particularly to the fragments of the hydroxamic (C=O, NH) and phenolic functional groups. The position of the Amide I (C=O) band at 1 615 cm^{-1} suggests that hydrogen bonding occurs between the carbonyl and hydroxy groups. The Amide II (ν(C—N) and δ(N—H)) band lies at 1 575 cm^{-1}, the ν(C—O) vibration of the phenolic functional group gives rise to a strong band at 1 355 cm^{-1} (ref.[19]).

The purity of the substance, as indicated by the results of the neutralization potentiometric titrations, is sufficiently high (99·5%). The absorption curves of H_2L (240—350 nm) measured at pH 3·50—11·50 exhibit a band at 320 nm which belongs to the L^{2-} and HL^- species; the undissociated H_2L species gives rise to a weak band at 295 nm. The wavelengths of 320 and 325 nm were used for the evaluation of the dissociation constants (the absorption maximum at 295 nm is unsuitable because of the low absorbance difference for H_2L and HL^-). The curve of the $A = f(pH)$ dependence and the distribution diagram of the various species is shown in Fig. 1. Graphical analysis (Fig. 2) gives evidence that at pH 6·74—8·15 as well as pH 9·20 to 10·00, always one proton is detached during the dissociation of the acid. The relations and transformations used for the determination of the molar absorptivities, number of protons eliminated during the reaction and dissociation constants were derived from ref.[20] The dissociation constants calculated from the curves of the dependence $A = f(pH)$ are given in Table I. We assume that pK_{a1} and pK_{a2} are associated with the deprotonation of the hydroxamic group and the phenolic group, respectively.

Like other hydroxamic acids, salicylhydroxamic acid forms coloured complexes with a number of metal ions in strongly or weakly acidic solutions; the most intense

are the colours with Fe(III) (red-violet), V(V) (red-violet), Cu(II) (blue), Ni(II) (green), Mo(VI) (lemon yellow), and UO_2^{2+} (orange). Colourless complexes are obtained with some metals such as Pb, Cd, Zn or Hg.

The complex formation between salicylhydroxamic acid and iron(III) was studied spectrophotometrically at pH 0–2 (in less acidic solutions a low-soluble complex

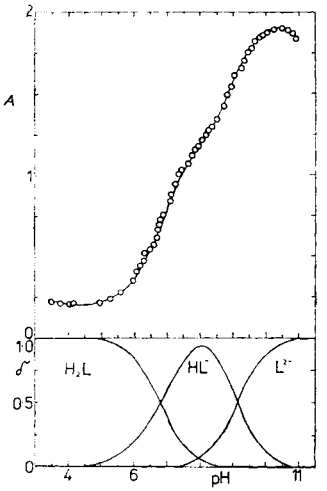

FIG. 1

Dependence $A = f(\mathrm{pH})$ and distribution diagram of salicylhydroxamic acid; $\lambda =$ $= 320$ nm, $l = 3{\cdot}48$ cm, $c_L = 80$ μmol l^{-1}, $I = 1{\cdot}0$ ($NaClO_4$)

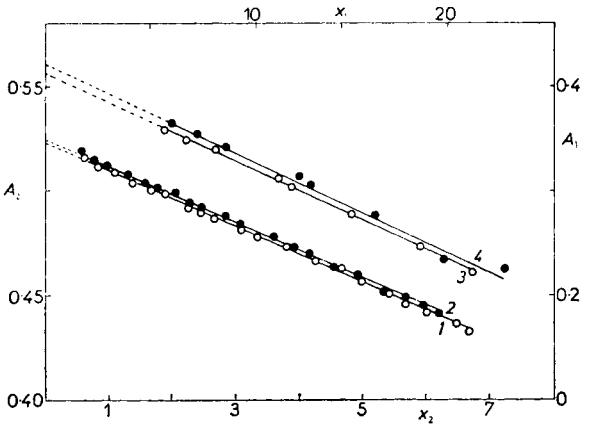

FIG. 2

Graphical analysis of the $A = f(\mathrm{pH})$ dependence of salicylhydroxamic acid; $c_L = 80$ μmol l^{-1}, $l = 1$ cm. λ(nm), pH: *1* 325, 6·74–8·15; *2* 320, 6·74–8·15; *3* 325, 9·20–10·00; *4* 320, 9·20–10·00. Symbols: $x_1 = (A - \varepsilon_{H_2L} c_L)\,[\mathrm{H}] \cdot 10^8$ (scale A_1), $x_2 = (A - \varepsilon_{HL} c_L)\,[\mathrm{H}] \cdot 10^{11}$ (scale A_2)

precipitates) and in conditions of excess reagent or iron (Fig. 3). The absorbances increase with increasing pH over the entire wavelength region, indicating the formation of a single complex with the absorption maximum at 530–535 mn. The molar absorptivities were obtained by graphical analysis of the $A = f(\mathrm{pH})$ curves. This analysis confirmed that the complex formation in strongly acidic solutions is asso-

TABLE I

Dissociation constants of salicylhydroxamic and 4-hydroxybenzohydroxamic acids and constants of formation of their iron(III) complexes at 20°C, $I = 1{\cdot}0$ ($NaClO_4$)

Equilibrium	Constant	log K			
		salicylhydroxamic acid		4-hydroxybenzo-hydroxamic acid	
		a	*b*	*a*	*b*
$[HL][H]/[H_2L]$	K_{a1}	$6{\cdot}73 \pm 0{\cdot}04$	$6{\cdot}83 \pm 0{\cdot}02$	$7{\cdot}65 \pm 0{\cdot}04$	$7{\cdot}77 \pm 0{\cdot}03$
$[L][H]/[HL]$	K_{a2}	$9{\cdot}15 \pm 0{\cdot}02$	$9{\cdot}15 \pm 0{\cdot}04$	—	—
$[FeHL][H]/[Fe][H_2L]$	$^*K$	—	$2{\cdot}36 \pm 0{\cdot}04$	—	$2{\cdot}33 \pm 0{\cdot}03$
$[FeHL]/[Fe][HL]$	β_1	—	$9{\cdot}09 \pm 0{\cdot}02$	—	$9{\cdot}98 \pm 0{\cdot}03$

[a] Potentiometric determination; [b] spectrophotometric determination.

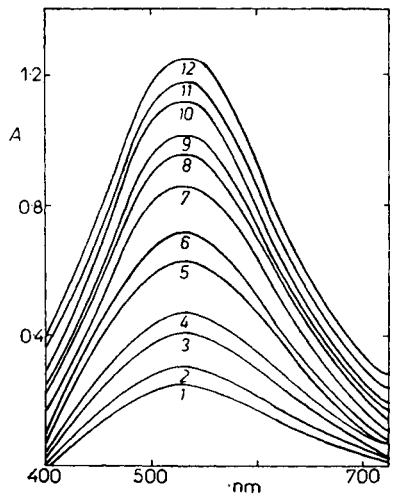

FIG. 3

Absorption curves of Fe(III)-salicylhydroxamic acid complex; $c_L = 0{\cdot}25$ mmol l^{-1}, $c_{Fe} = 1{\cdot}25$ mmol l^{-1}, $I = 1{\cdot}0$ ($NaClO_4$), $l = 3{\cdot}48$ cm. pH: *1* 0·024, *2* 0·127, *3* 0·269, *4* 0·367, *5* 0·492, *6* 0·682, *7* 0·878, *8* 1·029, *9* 1·132, *10* 1·359, *11* 1·472 *12* 1·628

ciated with the elimination of one proton,

$$Fe^{3+} + H_2L \rightleftarrows FeHL^{2+} + H^+ . \qquad (A)$$

For examining whether the phenolic group takes part in the chelation, 4-hydroxybenzohydroxamic acid was also synthesized and studied in the same manner as salicylhydroxamic acid (Table II). The composition and formation constant of the iron(III) complex of this reagent were found to approach closely those for the iron(III) complex of salicylhydroxamic acid (Table I). From this it can be inferred that in the hydrogencomplex $FeHL^{2+}$, only carbohydroxamic functional group takes part in the coordination of iron, one proton being split off and the five-membered chelate ring formed. It should be noted that this complex exists both in solutions with a concentration excess of iron(III) and in solutions with an excess of reagent.

The preparation and properties of the low-soluble complex forming in the region of pH 2–8·5 were also investigated. The results of elemental analysis confirm its composition as $Fe(OH)(HL)_2.3\,H_2O$. The IR spectrum of this complex suggests that iron in it is coordinated *via* the oxygen atoms of the carbonyl group and the deprotonated hydroxamic OH group: the Amide I and Amide II bands, lying at 1 600 and 1 656 cm^{-1}, are 15 and 10 cm^{-1}, respectively, shifted to lower wavenumbers as compared with those of salicylhydroxamic acid itself; the shift of the ν(C—O) band of the phenolic group to 1 380 cm^{-1} can be explained in terms of an interaction of the oxygen atom of the protonated phenolic group with the central Fe(III) atom.

Salicylhydroxamic acid is well suited as a chelometric indicator for the determination of iron, exhibiting a sharp colour change in the equivalence point. The optimum acidity is pH 1·2–1·5. Titrations at 20–50°C showed that the measurements are best performed at room temperature. The effect of the concentration of iron on the accuracy was examined, and it was found that under the conditions given above,

TABLE II

Spectrophotometric characteristics of salicylhydroxamic acid and its iron(III) complex

Species	Region of existence pH	λ_{max} nm	ε l mol^{-1} cm^{-1}
H_2L	2–5	295	3 200
HL^-	7·5–8·5	320	4 290
L^{2-}	10·5–11·5	320	6 840
$FeHL^{2+}$	0·1–1·9	530	1 570

samples with 2 – 10 mg of iron can be analyzed, but even as little as 0·5 mg of iron in the volume used can be determined with $2 . 10^{-3}$M-Chelaton 3. Of the cations followed, Cd, Cr, Ca, Mn, Mg, Al, Zn, and Co do not interfere even in a tenfold excess over iron (at $\varrho_{Fe(III)} = 55{\cdot}85\ mg\ l^{-1}$); Pb and Bi do not interfere if present in the same amount as iron. NO_3^-, SO_4^{2-}, Cl^-, and $H_2PO_4^-$ anions do not interfere in a hundredfold excess.

In comparison with the conventional indicator, sulphosalicylic acid, salicylhydroxamic acidic is advantageous in that it can be employed in strongly acid solutions. The chelometric determination of iron in drugs using salicylhydroxamic acid as the indicator has been studied separately[21].

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