J. Chem. Eng. Data **2007,** *52,* 1571−1579

Equilibrium Studies of Binary Systems Involving Lanthanide and Actinide Metal Ions and Some Selected Aliphatic and Aromatic Monohydroxamic Acids

Mohamed M. Khalil,* Mohamed M. El-Deeb, and Rehab K. Mahmoud

Department of Chemistry, Faculty of Science, Beni-Suef University, Beni-Suef, Egypt

Formation of 1:1, 1:2, and 1:3 binary complexes of Th(IV), UO_2(II), Ce(III), and La(III) metal ions with acetohydroxamic acid (Aha), benzohydroxamic acid (Bha), and salicylhydroxamic acid (Sham) was investigated in aqueous medium using the potentiometric technique at 25 °C and $I = 0.10$ mol·dm^{-3} $NaNO_3$. The order of stability of the complexes was investigated and is discussed in terms of both the nature of hydroxamic acid and the metal ion. The concentration distribution of various species formed in solution was evaluated. Evaluation of the effect of temperature of the medium on the ionization process of ligands and the stability of 1:1 binary complexes has been performed at $I = 0.10$ mol·dm^{-3} $NaNO_3$, and the corresponding thermodynamic parameters have been calculated and are discussed as well. Evaluation of the effect of ionic strength of the medium on both the dissociation process and the stability of metal complexes for Sham binary systems is reported. In addition, the dissociation constants of Aha and stability constants of its 1:1 binary complexes were determined in various water + dioxane mixtures under the experimental conditions ($t = 25$ °C, $I = 0.10$ mol·dm^{-3} $NaNO_3$). It was concluded that solvent effects have a profound influence on the proton−ligand and metal−ligand stability constants. Confirmation of the formation of binary complexes in solution has been done using conductivity, cyclic voltammetry, and UV−visible spectroscopic measurements.

Introduction

Over the past few decades, the coordination chemistry of actinides has been extensively investigated.[1−6] Actinides are present in sediments and in ground and surface waters as a result of both natural processes and anthropogenic activities.[7] They have remarkable environmental impact and are dangerous for human health because, like all heavy metals, they have high chemical toxicity and great affinity toward biological systems containing phosphoric, sulfuric, and oxygenated groups.[8−13] The necessity of treating and disposing large amounts of nuclear wastes in a safe and economical manner has generated a significant interest in recent years in the coordination chemistry of actinides in solution, especially at elevated temperatures. It is well-known that the temperature of the waste in storage tanks can be up to 90 °C,[14] whereas that in the vicinity of the waste forms in the repository can be as high as 100 °C to 300 °C.[15,16] Thus, complexation of actinides with ligands in solution at elevated temperatures should be understood. The majority of the literature data are obtained at or near 25 °C.[17] It is noted that lanthanide ions, which are able to hydrolyze nucleic acids, can form a variety of coordination compounds in aqueous solution. These complexes can be conjugated to RNA or DNA oligomers that bind to a specific site in the target nucleic acid.[18−23] Site-specific artificial enzymes can be considered as essential tools for biotechnology in the future.[24] Despite the existence of a very large amount of literature data on the solution chemistry of the actinide and lanthanide ions, interactions of these ions in aqueous media with molecules of biological interest, such as hydroxamic acids, have been little investigated. Since the discovery of hydroxamic acids by Wahlroos and Virtanen[25] in 1959, and over the past decades, their chemistry and biochemistry have attracted considerable attention, due to their pharmacological, toxicological, and pathological properties. It is well-known from the literature[26−28] that the unsubstituted aliphatic hydroxamic acids, for example, acetohydroxamic acid (Aha), are well established as effective urease inhibitors in plants[28] and have been shown to effectively inhibit ureolytic activity and/or to lower blood ammonia levels in sheep,[26] dogs, and man.[27] The potential applications of these compounds, in the treatment of hepatic coma and in the improvement of nitrogen utilization by ruminant animals, have led to the present series of their physiological disposition in the animal body.

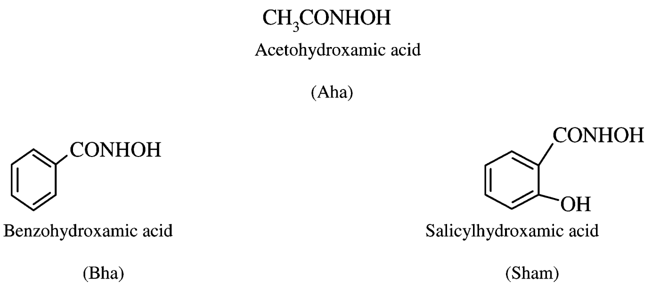
Recently, a new oral iron chelator, salicylhydroxamic acid (Sham), has been developed and found to have a promising advantage, because no toxicity has, as yet, been recorded.[29] The iron chelating property is due to the presence of the hydroxamic acid moiety (−CO−NOH), which it shares with desferrioxamine B and another lower molecular weight iron chelator, acetohydroxamic acid, which may also have a potential ability as an iron chelator.[30] In Egypt, it is recommended to use salicylhydroxamic acid as an effective and low-cost pharmaceutical drug. Sham, being a hydroxamic acid derivative of salicylic acid, exerts dual urolithiatic activity. It inhibits both the urease enzyme activity and sulfation of mucopolysaccharides, which are responsible for renal stone formation. The design and synthesis of ligands for biomedical applications in fields such as anticancer applications have become of great importance. One of those ligands is the hydroxamate molecule. Hydroxamic acids have been found to react with both proteins and nucleic acids.[31] Anticancer properties of the three hydroxamic acids, taken into consideration in this paper, [Aha, benzohydroxamic acid (Bha), and Sham] have been investigated.[32] Therefore, the present study focuses on the detailed complexation behavior of actinide and lanthanide metal ions with the above-mentioned hydroxamic acids.

---

* Corresponding author. E-mail: magdy_mmagdy@yahoo.com.

10.1021/je600541a CCC: $37.00

Scheme 1



Stabilities of binary complexes of both transition and alkaline earth metal ions with salicylhydroxamic acid have been recently investigated, in our laboratory, using the potentiometric technique at 25 °C and $I = 0.10$ mol·dm^{-3} $NaNO_3$.[33] More recently, the formation of binary complexes of the divalent [Cu(II), Co(II), Ni(II), and Zn(II)] and trivalent [Fe(III), Al(III), and Cr(III)] metal ions with the three hydroxamic acids was also studied by us[34] potentiometrically under the same experimental conditions and also at four different temperatures ranging from 25 °C to 55 °C for investigating the dissociation process of Bha as well as its 1:1 binary complexes with the trivalent metal ions mentioned above. As our research program aims to study the interaction of metal ions with biologically relevant ligands,[33−36] the present paper focuses on the detailed complexation behavior of actinide and lanthanide metal ions with the considered hydroxamic acids.

Experimental Section

Materials and Solutions. Aha and Bha were Sigma products. Sham was purchased in a pure form from the Nasr Pharmaceutical Chemicals Co., Egypt. The purity of hydroxamic acids and the concentrations of the stock solutions were determined by Gran's method.[37] The metal salts were provided by BDH as nitrates. Stock solutions of the metal salts were prepared in bidistilled water, and the metal concentration was standardized with the disodium salt of EDTA.[38] Carbonate-free sodium hydroxide (titrant, prepared in 0.10 mol·dm^{-3} $NaNO_3$ solution) was standardized potentiometrically with KH phthalate (Merck AG). A nitric acid solution (≈0.04 mol·dm^{-3}) was prepared and used after standardization. Sodium hydroxide, nitric acid, and sodium nitrate were from Merck p.a. Dioxane was of high purity (spectro grade product).

Apparatus and Procedure. Potentiometric pH measurements were performed using a Metrohm 702 titroprocessor equipped with a 665 dosimat (Switzerland). The precision of the instrument was (±0.001) pH unit. The pH titrations were carried out in an 80 cm^3 commercial double-walled glass vessel. The ionic strength of the solutions was maintained at a constant level by using the desired concentration of $NaNO_3$ solution as supporting electrolyte, and temperature was adjusted inside the cell at the desired temperature, by circulating thermostated water using an oil-thermostated setup. A computer program (GLEE, glass electrode evaluation)[39] has been used for the calibration of the glass electrode, in terms of $pH = -\log [H^+]$, by titrating HNO_3 (at the same temperature and ionic strength or for solutions containing different v/v % dioxane) against a NaOH standard solution. The resulting titration data were used to calculate the standard electrode potential $E°$ and K_w for water before each experiment. The investigated solutions were prepared (total volume = 50 cm^3) and titrated potentiometrically against standard CO_2-free NaOH (0.10 mol·dm^{-3}) solution. A stream of nitrogen was passed throughout the course of the experiment to exclude the adverse effect of atmospheric carbon dioxide.

The ligand concentrations were varied in the range of $1 \cdot 10^{-3}$ mol·dm^{-3} to $6 \cdot 10^{-3}$ mol·dm^{-3}. Three to four different metal-to-ligand ratios, ranging from 1:1 to 1:4, were used. The initial estimates of the ionization constants of the ligands as well as the stability constants of metal−ligand binary complexes were calculated by adopting the Irving and Rossotti technique.[40,41] The equilibrium constants were refined with a computer program based on an unweighted linear least-squares fit. The stoichiometries and stability constants of the complex species formed in solution were determined by examining various possible composition models for the systems studied. The model was that which gave the best statistical fit and plausibly consists of the 1:1 and 1:2 complex species for all systems investigated. In addition, the 1:3 complex for the binary systems Th(IV)−Sham, Ce(III)−Bha, Ce(III)−Aha, and La(III)−Aha has been detected.

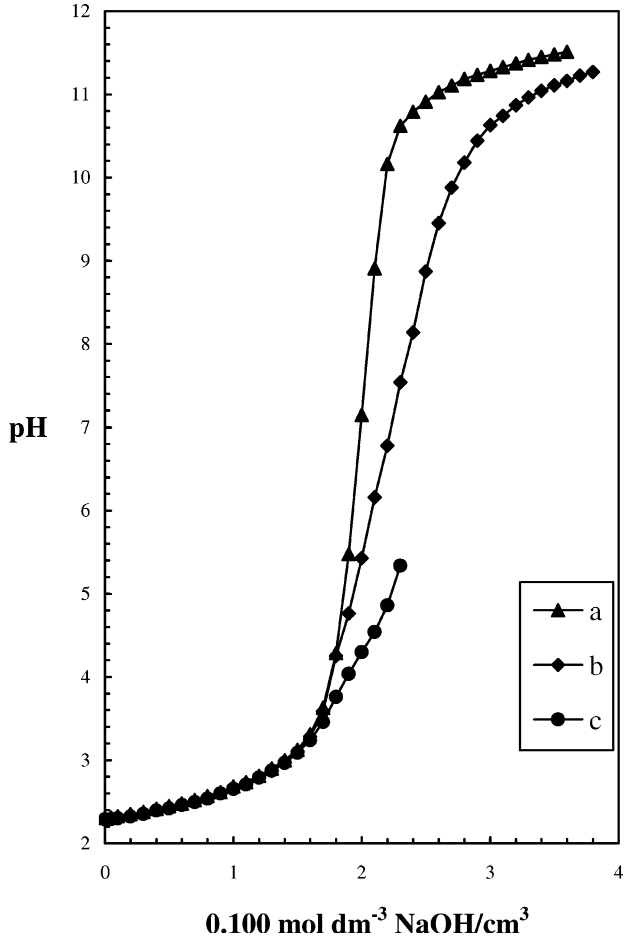


Figure 1. Potentiometric pH titration curves for UO_2(II)−Sham system at 25 °C and $I = 0.10$ mol·dm^{-3} $NaNO_3$: (a) $4 \cdot 10^{-3}$ mol·dm^{-3} HNO_3; (b) solution a + $1 \cdot 10^{-3}$ mol·dm^{-3} Sham; (c) solution b + $2.5 \cdot 10^{-4}$ mol·dm^{-3} UO_2(II).

Accounting for the differences in acidity, basicity, dielectric constant, and ion activities for solutions containing different proportions (5 % to 50 %) v/v of dioxane relative to pure aqueous ones, the pH values of the former solutions were corrected in accordance with the method described by Douheret.[42]

Each of the investigated solutions was thermostated at the required temperature with an accuracy of ±0.10 °C, and the solutions were left to stand at this temperature for about 15 min before titration. Magnetic stirring was used during all titrations. About 100 to 140 experimental data points were available for evaluation in each system. The titration was repeated at least five times for each titration curve. All calculations were made with the aid of a computer program based on an unweighted linear least-squares fit.

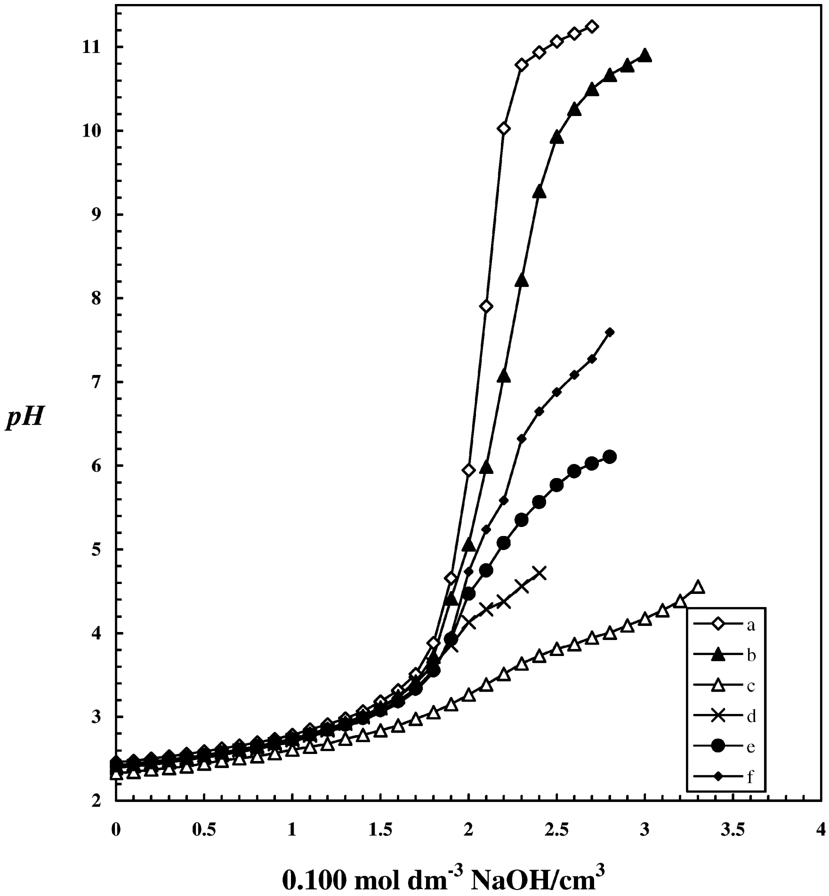


Figure 2. Potentiometric pH titration curves for Aha and its metal complexes at 25 °C and $I = 0.10$ mol·dm^{-3} $NaNO_3$: (a) $4 \cdot 10^{-3}$ mol·dm^{-3} HNO_3; (b) solution a + $1 \cdot 10^{-3}$ mol·dm^{-3} Aha; (c) solution b + $1 \cdot 10^{-3}$ mol·dm^{-3} Th(IV); (d) solution b + $1 \cdot 10^{-3}$ mol·dm^{-3} UO_2(II); (e) solution b + $1 \cdot 10^{-3}$ mol·dm^{-3} Ce(III); (f) solution b + $1 \cdot 10^{-3}$ mol·dm^{-3} La(III).

Table 1. Dissociation Constants (± SD) of the Ligands and Stability Constants (± SD) of Their 1:1, 1:2, and 1:3 Binary Complexes at 25 °C and $I = 0.10$ mol·dm^{-3} $NaNO_3$

	Aha pK_{a1} = 9.40 ± 0.01			Bha pK_{a1} = 8.63 ± 0.03			Sham pK_{a1} = 7.40 ± 0.02 pK_{a2} = 9.78 ± 0.01		
cation	log k_1	log k_2	log k_3	log k_1	log k_2	log k_3	log k_1	log k_2	log k_3
Th(IV)	10.50 ± 0.03	9.20 ± 0.02		8.97 ± 0.02	8.56 ± 0.03		14.40 ± 0.01	12.60 ± 0.01	9.15 ± 0.02
UO_2(II)	8.55 ± 0.02	7.94 ± 0.04		7.42 ± 0.03	6.70 ± 0.01		11.90 ± 0.02	9.70 ± 0.02	
Ce(III)	7.85 ± 0.01	6.90 ± 0.04	5.76 ± 0.01	7.25 ± 0.05	6.40 ± 0.04	5.22 ± 0.02	11.00 ± 0.01	8.69 ± 0.01	
La(III)	6.99 ± 0.02	6.05 ± 0.03	3.98 ± 0.02	6.75 ± 0.02	5.90 ± 0.02		10.90 ± 0.03	7.77 ± 0.02	

Conductometric Measurements. Conductometric titrations were followed with a HANNA conductivity meter HI-98304. The following mixture was titrated conductometrically against 0.10 mol·dm^{-3} NaOH solution: $1 \cdot 10^{-2}$ mol·dm^{-3} La(III) (20 cm^3) + $1 \cdot 10^{-2}$ mol·dm^{-3} hydroxamic acid (20 cm^3).

Cyclic Voltammetry Measurements. Cyclic voltammetric measurements were collected using potentiostat/galvanostate wenking PGS 95 with a single-compartment voltammetric cell equipped with a platinum working electrode (area = 0.50 cm^2), a Pt wire counter electrode, and a SCE as a reference electrode. In a typical experiment, a sample volume of 25 cm^3 containing the free metal ion $1 \cdot 10^{-3}$ mol·dm^{-3} Th(IV) (a), $1 \cdot 10^{-3}$ mol·dm^{-3} Th(IV) + $1 \cdot 10^{-3}$ mol·dm^{-3} hydroxamic acid (b) was used. The ionic strength of the studied solutions was adjusted at 0.10 mol·dm^{-3} using a $NaNO_3$ solution. All solutions were investigated in water at 25 °C. The solutions were purged with nitrogen for 120 s, and then the potential was scanned at the scan rate 25 mV·s^{-1} from (+0.10 to −0.10) V.

Spectrophotometric Measurements. The absorption spectra of solutions to be tested were recorded on a Shimadzu Corp. UV-1601 (PC)S spectrophotometer in the range of 300 nm to 500 nm using 1 cm matched quartz cells.

Results and Discussion

The formulas of the hydroxamic acids investigated are shown in Scheme 1. The proton dissociation constants of the hydroxamic acids studied have been redetermined at 25 °C and $I = 0.10$ mol·dm^{-3} $NaNO_3$ to obtain values using the same experimental procedures as used in the study of binary systems and are in good agreement with data found in the literature.[33,43,44] The differences in pK_a values for Aha, Bha, and Sham have been discussed previously by us.[34]

Representative potentiometric pH titration curves for the UO_2(II)−Sham binary system are shown in Figure 1. Analysis of the complexed ligand curves indicates that the addition of metal ion to the free ligand solution shifts the buffer region of the ligand to lower pH values. This shows that complex formation reaction proceeds by releasing protons from such ligands. The stability constants of 1:1, 1:2, and 1:3 binary complexes of the considered ligands have been determined at

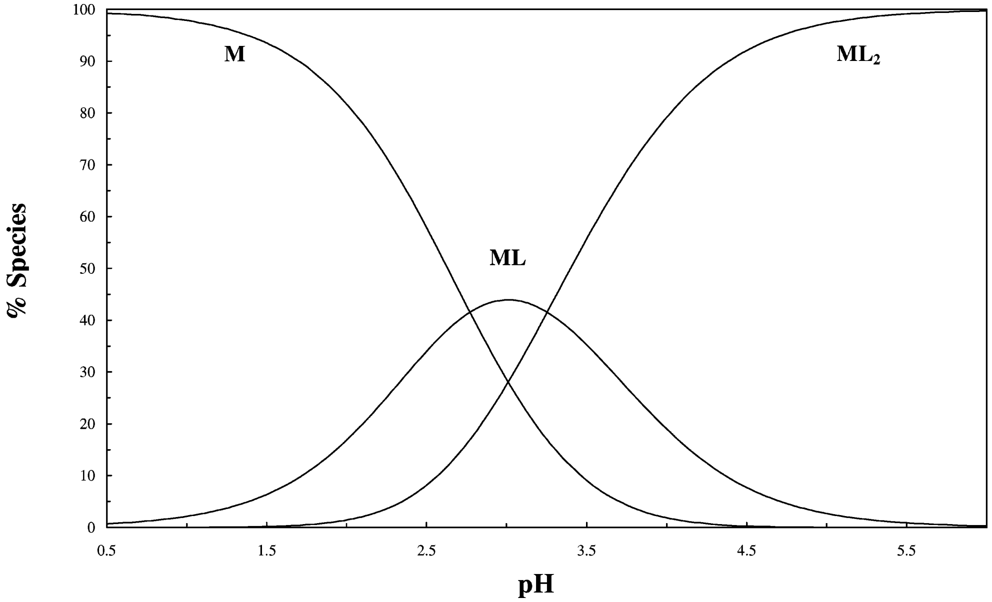


Figure 3. Distribution diagram for the system Th(IV)−Bha at $I = 0.10$ mol·dm^{-3} NaNO$_3$ and $t = 25$ °C. Concentration: $C_{\mathrm{Th(IV)}} = 3 \cdot 10^{-4}$ mol·dm^{-3}; $C_{\mathrm{Bha}} = 1 \cdot 10^{-3}$ mol·dm^{-3}.

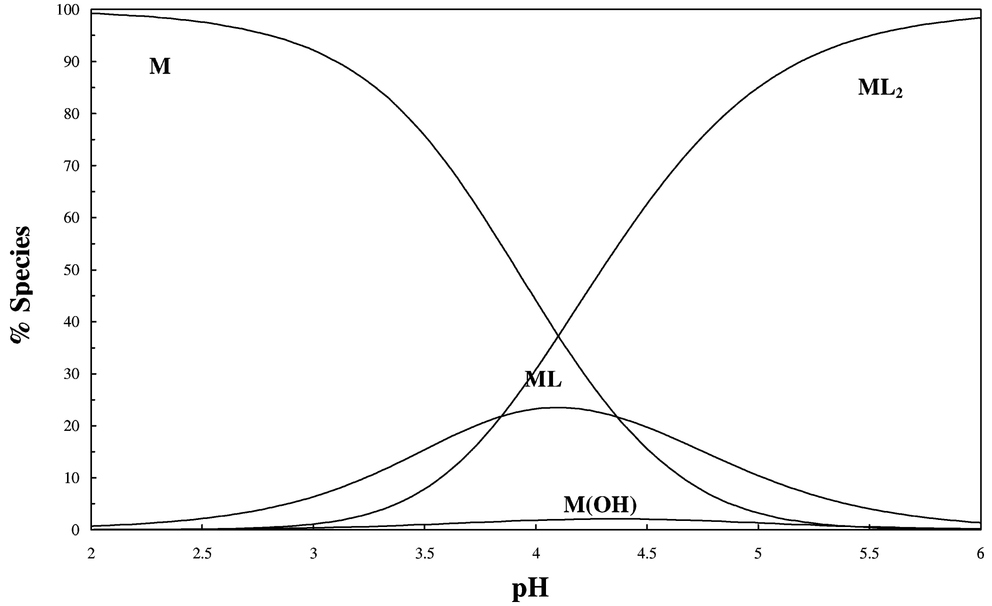


Figure 4. Distribution diagram for the system UO$_2$(II)−Aha at $I = 0.10$ mol·dm^{-3} NaNO$_3$ and $t = 25$ °C. Concentration: $C_{\mathrm{UO_2(II)}} = 3 \cdot 10^{-4}$ mol·dm^{-3}; $C_{\mathrm{Aha}} = 1 \cdot 10^{-3}$ mol·dm^{-3}.

25 °C and $I = 0.10$ mol·dm^{-3} NaNO$_3$. Figure 2 displays a representative set of experimental titration curves for Aha and its 1:1 binary complexes involving the investigated metal ions.

The equilibria involved in the formation of the metal−hydroxamic acid binary complexes may be represented as follows:

$$\mathrm{M} + \mathrm{L} \rightleftharpoons \mathrm{ML} \qquad K_1 = \frac{[\mathrm{ML}]}{[\mathrm{M}][\mathrm{L}]} \tag{1}$$

$$\mathrm{ML} + \mathrm{L} \rightleftharpoons \mathrm{ML}_2 \qquad K_2 = \frac{[\mathrm{ML}_2]}{[\mathrm{ML}][\mathrm{L}]} \tag{2}$$

$$\mathrm{ML}_2 + \mathrm{L} \rightleftharpoons \mathrm{ML}_3 \qquad K_3 = \frac{[\mathrm{ML}_3]}{[\mathrm{ML}_2][\mathrm{L}]} \tag{3}$$

Examination of the stability constants of the binary complexes investigated (Table 1) leads to the following remarks:

(a) The observed order of stability of binary systems with respect to the ligand hydroxamic acid is Sham > Aha > Bha.[34]

(b) The complex stability of the binary complexes with respect to the metal ion present follows the order Th(IV) > UO$_2$(II) > Ce(III) > La(III).[45,46]

(c) Also, it can be observed that the stability constants of the different 1:3 metal−ligand complexes are lower than those of the corresponding 1:2 complex species, and the stability constants of the latter complexes are lower than those of the corresponding 1:1 systems, as expected from statistical consideration. The Δlog k (log k_3 − log k_2 or log k_2 − log k_1) values are negative (Table 1). The reduction in the values of stepwise constants is principally due to the fact that the entropy contribution to the free energy change becomes less favorable from one step to the next (eqs 1 to 3).

Estimation of the concentration distribution of various complex species in solution provides a useful picture of metal ion binding in biological systems. The speciation diagrams

Table 2. Thermodynamic Quantities Associated with the Dissociation of the Ligands and the Interaction of Metal Ions with the Ligands at 1:1 Molar Ratio and $I = 0.10$ mol·dm^{-3} NaNO$_3$

		pK_a or log k_1			
ligand or complex	cation	$t = 25$ °C	$t = 35$ °C	$t = 45$ °C	$t = 55$ °C
Aha	H	9.40 ± 0.01	9.21 ± 0.04	9.13 ± 0.01	8.88 ± 0.02
1:1 binary	Th(IV)	10.50 ± 0.03	10.22 ± 0.01	10.07 ± 0.01	9.93 ± 0.01
complex	UO$_2$(II)	8.55 ± 0.02	8.38 ± 0.01	8.28 ± 0.02	8.20 ± 0.03
of Aha	Ce(III)	7.85 ± 0.01	7.76 ± 0.04	7.70 ± 0.03	7.61 ± 0.01
	La(III)	6.99 ± 0.02	6.92 ± 0.02	6.88 ± 0.01	6.85 ± 0.04
Bha	H$^{(b)}$	8.63 ± 0.04	8.58 ± 0.04	8.55 ± 0.02	8.50 ± 0.04
1:1 binary	Th(IV)	8.97 ± 0.02	8.80 ± 0.01	8.68 ± 0.03	8.59 ± 0.01
complex	UO$_2$(II)	7.42 ± 0.03	7.29 ± 0.03	7.22 ± 0.02	7.18 ± 0.01
of Bha	Ce(III)	7.25 ± 0.05	7.15 ± 0.04	7.08 ± 0.01	7.04 ± 0.03
	La(III)	6.75 ± 0.02	6.70 ± 0.02	6.67 ± 0.03	6.64 ± 0.01
Sham	H	7.40 ± 0.02	6.85 ± 0.01	6.53 ± 0.02	6.32 ± 0.03
	H	9.78 ± 0.01	9.69 ± 0.02	9.63 ± 0.03	9.57 ± 0.01
1:1 binary	Th(IV)	14.40 ± 0.01	13.64 ± 0.02	13.25 ± 0.01	12.92 ± 0.03
complex	UO$_2$(II)	11.90 ± 0.02	11.65 ± 0.04	11.52 ± 0.01	11.43 ± 0.02
of Sham	Ce(III)	11.00 ± 0.01	10.83 ± 0.02	10.73 ± 0.04	10.66 ± 0.01
	La(III)	10.90 ± 0.03	10.73 ± 0.01	10.63 ± 0.02	10.55 ± 0.03

		$\Delta H°$	$\Delta G°$	$\Delta S°$
ligand or complex	cation	kJ·mol^{-1}	kJ·mol^{-1}	J·mol^{-1}·K^{-1}
Aha	H	28.76	53.65	−83.66
1:1 binary	Th(IV)	−37.88	−59.93	74.67
complex	UO$_2$(II)	−23.43	−48.80	85.01
of Aha	Ce(III)	−15.73	−44.81	97.65
	La(III)	−10.19	−39.90	99.57
Bha	H$^{(c)}$	94.53	48.63	−160.71
1:1 binary complex	Th(IV)	−27.92	−51.20	78.12
complex	UO$_2$(II)	−17.87	−42.35	82.33
of Bha	Ce(III)	−15.10	−41.38	88.08
	La(III)	−7.55	−38.53	103.97
Sham	H	96.16	42.24	−179.69
	H	17.42	55.82	−128.29
1:1 binary	Th(IV)	−109.08	−82.19	89.89
complex	UO$_2$(II)	−34.85	−67.92	111.05
of Sham	Ce(III)	−25.42	−62.76	124.46
	La(III)	−23.60	−62.21	128.29

obtained for Th(IV)−Bha and UO$_2$(II)−Aha systems are shown in Figures 3 and 4, respectively. In all of the species distributions, the concentration of the complex increases with increasing pH. The hydrolytic constant used for UO$_2$(II) was previously determined.[47] The concentration distribution of various complex species existing in solution as a function of pH was obtained using the SPECIES program.[48]

Stability constants of binary complex species ML, ML$_2$, and ML$_3$ were determined successfully at pH < 4.5.[1−6,24,49−51] The formation of monohydroxo complexes could not be detected; the formation constants of the corresponding complex species were rejected by our computer program. All side reactions due to hydrolysis of lanthanides or actinides have not been included in our calculations. The mononuclear hydrolytic species, for example, UO$_2$OH$^+$, is formed at low uranyl concentrations. The bi- and trinuclear hydrolytic species were found to be stabilized only in chloride media by the formation of a fairly stable ion pair (UO$_2$)$_3$(OH)$_4$Cl$^+$.[7,52] These hydrolytic species are generally formed with high percentages at pH ≥ 5, even in the presence of O−O donor ligands.

The effect of temperature of the medium on both the dissociation of hydroxamic acids investigated as well as the stability of their 1:1 binary complexes was also studied at $I = 0.10$ mol·dm^{-3} NaNO$_3$. The values of the protonation and binary (1:1) equilibrium constants for each hydroxamic acid ligand are found to be linearly dependent on the inverse of temperature, indicating negligible change in heat capacity for each of these protonation and complexation reactions.[53] The equilibrium constants have been evaluated at four different temperatures (from 25 °C to 55 °C), along with the thermodynamic quantities, and the values obtained are cited in Table 2. The values of $\Delta H°$ for the ionization of the ligands are found to be positive, indicating the endothermic nature of the deprotonation process. The positive values of the standard free energy change ($\Delta G°$) for the dissociation processes of the ligands denote that the processes are not spontaneous. Also, the negative values of $\Delta S°$ are pointing to increased ordering due to association. The values of the formation constants decrease with temperature because the formation reactions are exthothermic (Le Chatelier's principle). The values of enthalpy changes ($\Delta H°$) for the binary systems investigated are negative. However, the complex formation process is spontaneous in nature, as characterized by the negative $\Delta G°$ values. The values of $\Delta S°$ substantiate the suggestion that the different binary complexes are formed due to the coordination of the ligand anion to the metal cation. Furthermore, the positive values of $\Delta S°$ suggest also a desolvation of the ligands, resulting in weak solvent−ligand interactions, to the advantage of the metal ion−ligand interaction.[54]

The ionic strength of the medium is considered as a measure of total electrolyte concentration. It is related to the activity coefficients of the ions in solution by the following relationship:[54]

$$-\log f_i = 0.51 Z_i^2 \sqrt{I}/(1 + 0.33\alpha_i \sqrt{I})$$

The dissociation constants of Sham and the stability constants of its 1:1 binary complexes at 25 °C and different ionic strengths have been evaluated, and the values obtained are cited in Table 3. The plot of pK values for Sham and log k_1 for its binary systems versus $\sqrt{I}$ is linear as shown in Figure 5. The thermodynamic equilibrium constants, at $I = 0.0$ [p$K_{a1} = 6.99 \pm 0.02$, p$K_{a2} = 9.33 \pm 0.01$, and log $k_1 = 15.25 \pm 0.03$, 12.38 ± 0.02, 11.39 ± 0.03, and 11.20 ± 0.04 for Th(IV), UO$_2$(II), Ce(III), and La(III) binary systems, respectively] were determined by applying linear regression analysis.

To shed more light on both the protonation process of Aha and, consequently, its complexation process in various water + organic solvent mixtures, dioxane (an aprotic nonionizing coorganic solvent) was chosen. The observed increase in the protonation constants of Aha upon enrichment of the solvent with dioxane can be ascribed to a lowering of the dielectric

Table 3. Dissociation Constants of Sham and Stability Constants of Its 1:1 Binary Complexes at 25 °C and Different Ionic Strengths

			log $K^{\mathrm{M}}_{\mathrm{M(Sham)}}$			
ionic strength I (NaNO$_3$)	pK_{a1}	pK_{a2}	Th(IV)	UO$_2$(VI)	Ce(III)	La(III)
0.25	7.22 ± 0.03	9.53 ± 0.03	13.16 ± 0.02	11.25 ± 0.01	10.60 ± 0.01	10.19 ± 0.03
0.20	7.27 ± 0.02	9.65 ± 0.04	13.42 ± 0.06	11.54 ± 0.06	10.64 ± 0.02	10.34 ± 0.01
0.15	7.29 ± 0.05	9.70 ± 0.02	13.83 ± 0.03	11.77 ± 0.05	10.81 ± 0.04	10.53 ± 0.02
0.10	7.40 ± 0.02	9.78 ± 0.01	14.40 ± 0.01	11.90 ± 0.02	11.00 ± 0.01	10.90 ± 0.03
0.05	7.57 ± 0.06	9.92 ± 0.02	14.58 ± 0.04	12.05 ± 0.02	11.13 ± 0.03	10.98 ± 0.05
0.02	7.66 ± 0.02	10.03 ± 0.04	15.02 ± 0.02	12.17 ± 0.01	11.25 ± 0.02	11.04 ± 0.03

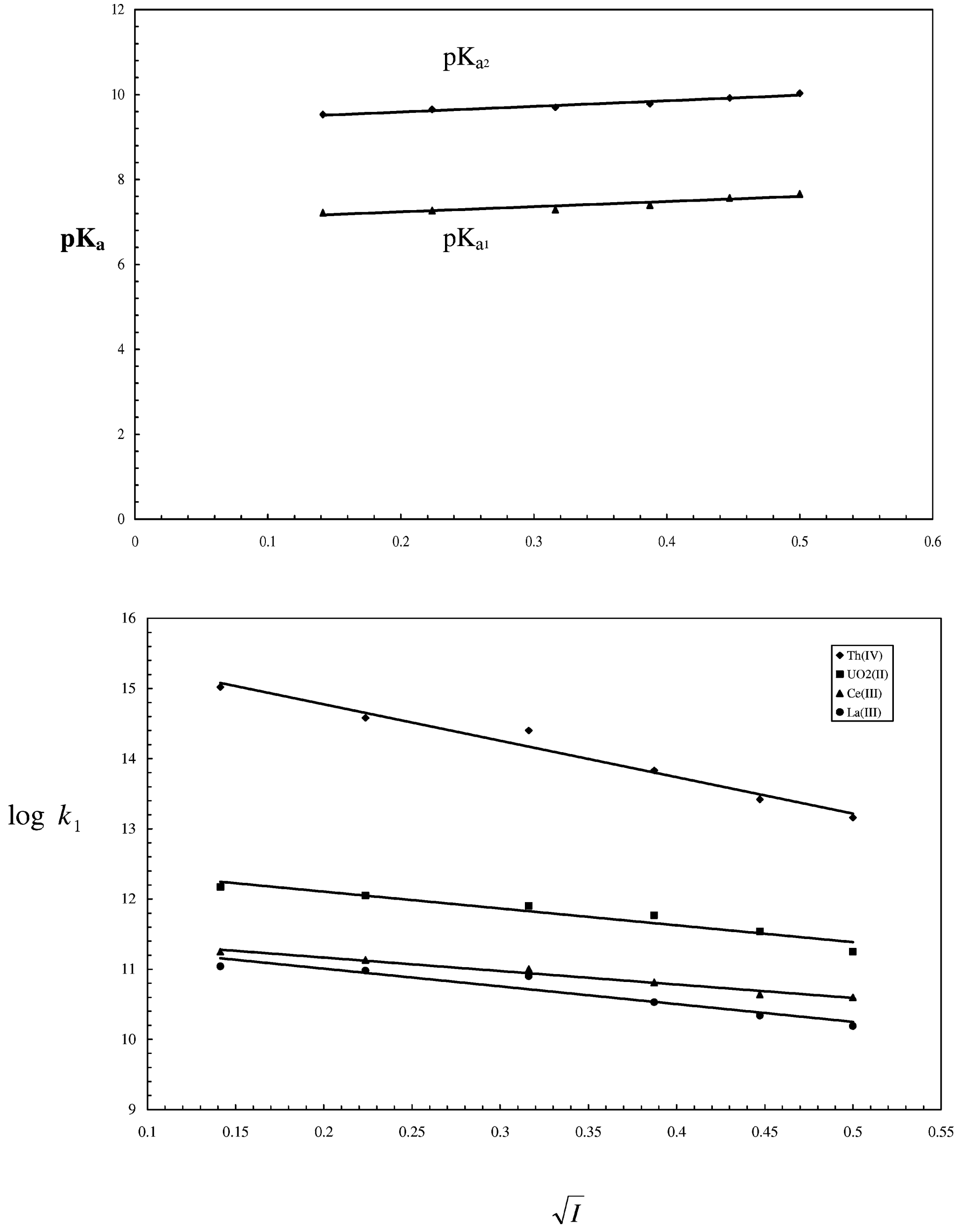


Figure 5. Plot of pK values of Sham and log k_1 for its binary systems versus $\sqrt{I}$ at 25 °C.

Table 4. Dissociation Constants of Aha and Stability Constants of Its 1:1 Binary Complexes in Aqueous Dioxane Media at 25 °C and I = 0.10 mol·dm^{-3} NaNO$_3$

			log k_1			
solvent composition % v/v	dielectric constant of the medium (ϵ)	pK_{a1}	Th(IV)	UO$_2$(II)	Ce(III)	La(III)
0.00	78.54	9.40 ± 0.01	10.50 ± 0.03	8.55 ± 0.02	7.85 ± 0.01	6.99 ± 0.02
5.00	77.71	9.44 ± 0.05	10.55 ± 0.03	9.19 ± 0.03	7.90 ± 0.02	7.23 ± 0.05
10.0	76.80	9.60 ± 0.03	10.61 ± 0.05	9.22 ± 0.05	8.00 ± 0.05	7.35 ± 0.03
20.0	74.74	9.66 ± 0.06	10.69 ± 0.04	9.26 ± 0.02	8.17 ± 0.02	7.42 ± 0.04
30.0	72.24	9.70 ± 0.02	10.72 ± 0.05	9.33 ± 0.04	8.27 ± 0.06	7.80 ± 0.02
40.0	69.19	9.72 ± 0.01	10.78 ± 0.01	9.50 ± 0.03	8.36 ± 0.02	8.17 ± 0.03

constant (ϵ = 77.71 and 65.32 for 5 % and 50 % v/v solvent composition, respectively, at 25 °C), which increases, in turn, the fraction of associated ions to form Bjerrum ion pairs and higher aggregates such as triple ions and dipole aggregates.[56,57] The concentration of free ions is very low, and the complexes are governed largely by ionic association reactions. The change in the relative permittivity of the medium (ϵ) influences the activity coefficient of the charged species, as reported previously by Coetzee and Ritchie.[58] Thus, the activity coefficient of the lanthanide or actinide ion and the hydroxamate anion will increase with an increasing amount of dioxane in the aqueous medium and, hence, the complex stability constants will increase. It is concluded that electrostatic effects, established from the change in the relative permittivity of the medium, play the major role in the magnitude of the complexes' stability, and the interactions between the metal ion and the ligand are predominately ionic. However, the variation of log k_1 values with the reciprocal of ϵ is not linear. This behavior shows that

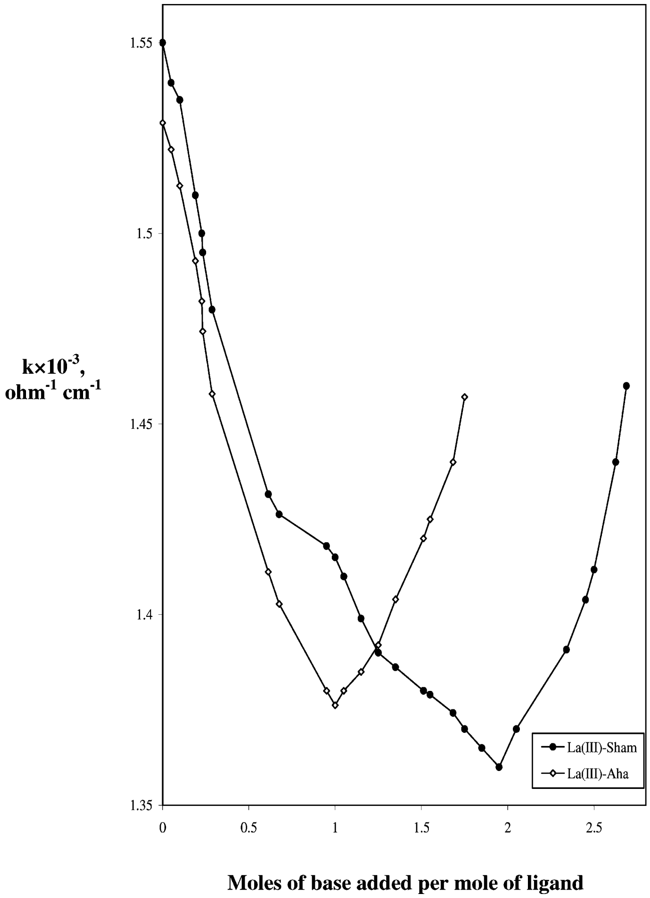


Figure 6. Conductometric titration curves for La(III)−Sham and La(III)−Aha binary systems.

the increase in the stability constant of the 1:1 binary complex, although mainly governed by the electrostatic effect, is influenced also by other solvent effects. Mui et al.[59] provided evidence suggesting the presence of the probable important interactions with solvent in the overall reactions of complex formation. Water molecules have a high tendency to develop H-bonds compared with other solvents.[60] Thus, the hydroxamate anion is less stabilized by hydrogen-bonding interactions as the amount of the organic solvent is increased in the medium, leading to an increase in the association of the anion with the positive metal ion, forming the metal complexes. The pK_a values of Aha and its stability constants for 1:1 complex species in aqueous dioxane media are cited in Table 4.

Conductometric titrations have been investigated to indicate the complexation behavior of the binary systems studied in solution. In Figure 6 representative conductometric titration curves for binary complexes of La(III) with Aha and Sham are displayed. The titration curves show an initial decrease and inflections at $a = 1$ and $a = 2$ for La(III)−Aha and La(III)−Sham systems, respectively (a = moles of base added per mole of ligand). This probably corresponds to the neutralization of H^+ ions resulting from the formation of the binary complexes. Beyond $a = 1$ for the La(III)−Aha titration curve and $a = 2$ for the La(III)−Sham curve, the conductance increases more uniformly due to the presence of an excess of NaOH.

Confirmation of the formation of the binary complexes in solution has been studied using the cyclic voltammetry technique. The binary systems of hydroxamic acids involving Th(IV) have been selected for study due to their high stability compared to the other binary systems studied as mentioned in the potentiometric studies. Figure 7 represents the cyclic voltammograms for the reduction of $1\cdot10^{-3}$ mol·dm^{-3} Th(IV) in the absence and presence of the hydroxamic acid ligands ($1\cdot10^{-3}$ mol·dm^{-3}) at $I = 0.10$ mol·dm^{-3} $NaNO_3$, 25 °C, and pH 3.5. The addition of hydroxamic acid ligand causes a slight shift of the cathodic peaks to more negative potential, indicating the formation of binary complexes in solution. From a comparison of the values of the reduction potentials of the three binary systems investigated, it is clear that the stability of complexes follows the order

$$\text{Th(IV)}-\text{Bha}\ (E_{\text{red}} = -70\ \text{mV}) < \text{Th(IV)}-\text{Aha}\ (E_{\text{red}} = -225\ \text{mV}) < \text{Th(IV)}-\text{Sham}\ (E_{\text{red}} = -380\ \text{mV})$$

This order is in accordance with the results obtained by potentiometric pH titrations under the same conditions as shown in Table 1.

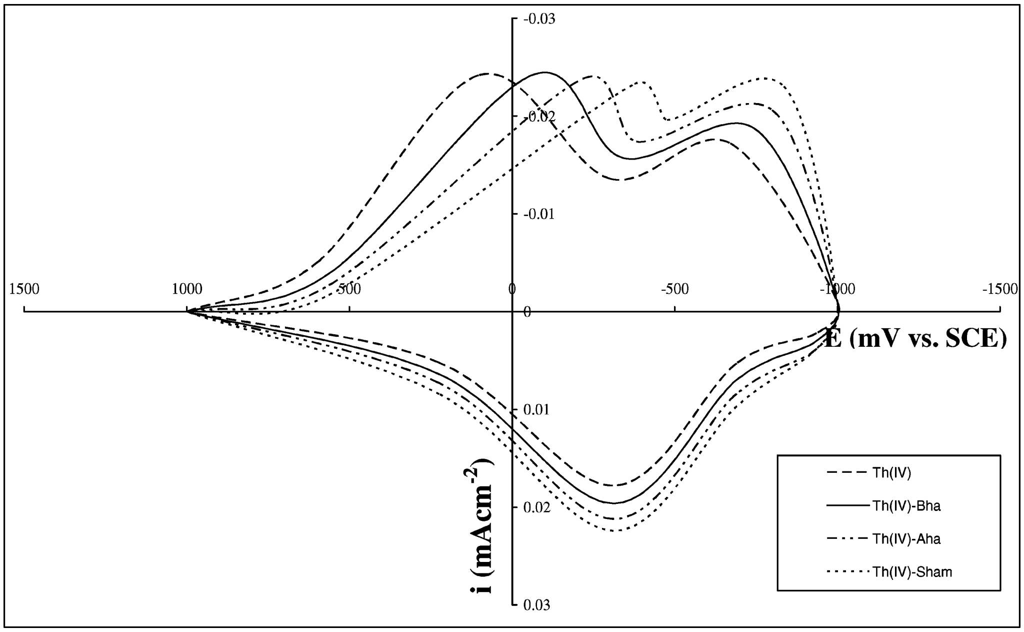


Figure 7. Cyclic voltammograms for the Th(IV) in the absence and presence of hydroxamic acid systems at $I = 0.10$ mol·dm^{-3} $NaNO_3$ and pH 3.5 with a scan rate of 25 mV·s^{-1} at 25 °C. $C_{\text{Th(IV)}} = 1\cdot10^{-3}$ mol·dm^{-3} and metal−ligand ratio of 1:1.

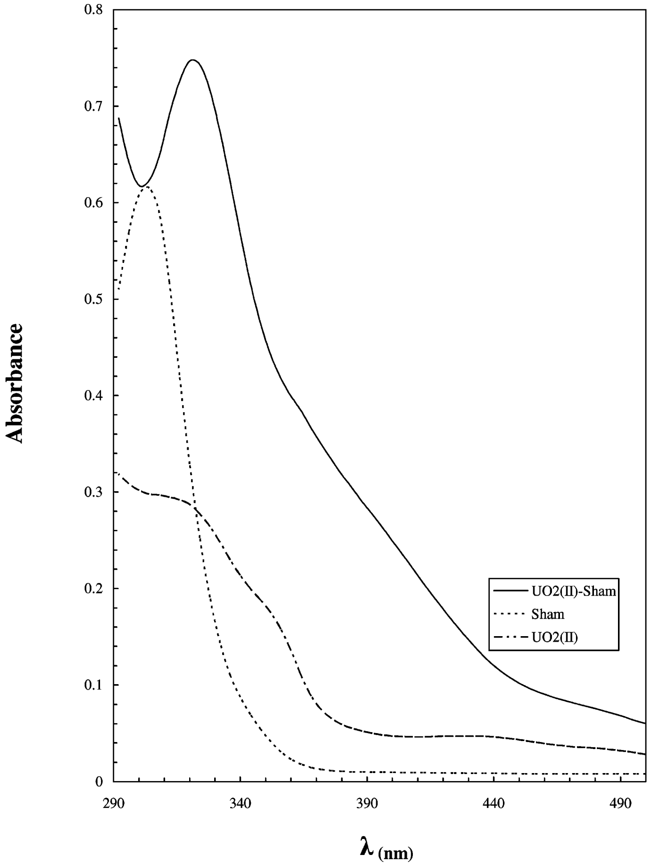


Figure 8. UV−visible absorption spectra for the UO_2(II)−Sham system at $I = 0.10$ mol·dm^{-3} $NaNO_3$, $C_{\text{Sham}} = 2 \cdot 10^{-4}$ mol·dm^{-3}, metal−ligand ratio of 1:1, and 25 °C.

The main features of the voltammograms are a four-electron oxidation step at negative potentials and a four-electron two reduction steps at negative potentials. The voltammetric behavior is shown to be quasireversible for all Th(IV)−hydroxamic acid binary systems investigated, and the irreversibility phenomenon increases according to the sequence

$$\text{Th(IV)−Sham} \rightarrow \text{Th(IV)−Aha} \rightarrow \text{Th(IV)−Bha}$$

The spectrum of a solution containing UO_2(II) ($2 \cdot 10^{-4}$ mol·dm^{-3}) exhibited no absorption when the solution was scanned versus a blank solution similarly prepared but containing no UO_2(II). A Sham solution ($2 \cdot 10^{-4}$ mol·dm^{-3}), scanned versus a blank solution similarly prepared but containing no Sham, shows an absorption band at about 304 nm. However, the solution containing Sham and UO_2(II) (1:1 molar ratio) undergoes a change in color, and the spectrum of the reaction mixture against a blank containing the same concentration of the ligand exhibits a new band at 316 nm. The latter is presumably due to the formation of a binary complex. The maximum color development for the binary system was attained at pH 6.20, and the absorbance decreases if the pH is increased further. The absorption spectra obtained are shown in Figure 8.

Acknowledgment

We express our sincere thanks to Jihan H. Ali, Assistant Lecturer, Department of English, Faculty of Arts, Beni-Suef University, for improving the scientific language of the paper.

Literature Cited

(1) Pettit, D.; Powell, K. *IUPAC Stability Constants Database*; Academic Software: Otley, U.K., 2001.

(2) May, P. M.; Murray, K. *Joint Expert Speciation System*; JESS Primer; Murdoch University: Murdoch, Australia, 2000.

(3) Martell, A. E.; Motokaitis, R. J.; Smith, R. M. *NIST-Database 46*, version 4; Gaithersburg, MD, 1997.

(4) Sillen, L. G.; Martell, A. E. *Stability Constants of Metal-Ion Complexes*; Supplement Special Publication 25; The Chemical Society: London, U.K., 1971, and references cited therein.

(5) (a) Martell, A. E.; Smith, R. M. *Critical Stability Constants*; Plenum Press: New York, 1977; Vol. 3. (b) Martell, A. E.; Smith, R. M. *Critical Stability Constants*, 2nd ed.; Plenum Press: New York, 1997; Vol. 6.

(6) Guillaumont, R.; Fanghänel, T.; Fuger, J.; Grenthe, I.; Neck, V.; Palmer, D. A.; Rand, M. H. In *Update on the Chemical Thermodynamics of Uuranium, Neptunium, Plutonium; Americium and Technetium*; Mompean, F., Illemassène, M., Domenechorti, C., Ben-Said, K. Eds.; Elsevier: Amsterdam, The Netherlands, 2003.

(7) Crea, F.; De Robertis, A.; De Stefano, C.; Sammartano, S. Dioxouranium(VI)−Carboxylate Complexes. A Calorimetric and Potentiometric Investigation of Interaction with Oxalate at Infinite Dilution and in NaCl Aqueous Solution at $I = 1.0$ mol L^{-1} and $t = 25$ °C. *Talanta* **2007**, *71* (2), 948−963.

(8) Ozmen, M.; Yurekli, M. Subacute Toxicity of Uranyl Acetate in Swiss-Albino Mice. *Environ. Toxicol. Pharmacol.* **1998**, *6* (2), 111−115.

(9) Domingo, J. L.; Llobet, J. M.; Thomas, J. M.; Corbella, J. Acute Toxicity of Uranium in Rats and Mice. *Bull. Environ. Contam. Toxicol.* **1987**, *39*, 168−174.

(10) Ballou, J. E.; Gies, R. A.; Case, A. C.; Haggard, D. L.; Buschbom, R. L.; Ryan, J. L. Deposition and Early Disposition of Inhaled $^{233}UO_2(NO_3)_2$ and $^{232}UO_2(NO_3)_2$ in the Rat. *Health Phys.* **1986**, *51* (6), 755−771.

(11) Harvey, R. B.; Kubena, L. F.; Lovering, S. L.; Mollenhauer, H. H.; Phillips, T. D. Acute Toxicity of Uranyl Nitrate to Growing Chicks: A Pathophysiologic Study. *Bull. Environ. Contam. Toxicol.* **1986**, *37* (1), 907−915.

(12) Nechay, B. R.; Thompson, J. D.; Palmer Saunders, J. Inhibition by Uranyl Nitrate of Adenosine Triphosphatases Derived from Animal and Human Tissues. *Toxicol. Appl. Pharmacol.* **1980**, *53* (3), 410−419.

(13) Bastide, J.; Bastide, P.; Malet, P.; Turchini, J. P.; Dastugue, G. Adenosine Triphosphatase Activity in the Kidneys of Rats and Mice During Experimental Poisoning with Uranyl Nitrate (Biochemical and Histochemical Study). *Therapie* **1966**, *21* (6), 1607−1616.

(14) Srinivasan, T. G.; Zanonato, P.; Di Bernardo, P.; Bismondo, A.; Rao, L. Complexation of Thorium(IV) with Malonate at Variable Temperatures. *J. Alloys Compds.* **2006**, *408−412*, 1252−1259.

(15) DOE. *Estimating the Cold War Mortgage: The 1995 Baseline Environmental Mangment Report*; DOE/EM-0232; U.S. Department of Energy: Wasington, DC, 1995.

(16) Cruse, J. M.; Gilchrist, R. L. Developing Waste Disposal Options in the Underground Storage Tank-Integrated Demonstration Program. In *Global 93, Future Nuclear Systems: Emerging Fuel Cycles and Waste Disposal Options*; American Nuclear Society: La Grange Park, IL, 1993; 1376 pp.

(17) Martell, A. E.; Smith, R. M. *Critical Stability Constants*; Plenum Press: New York, 1989; Vol. 6.

(18) Hettich, R.; Schneider, H. Supramolecular Chemistry. Part 71. [1]Evidence for Hydrolytic DNA Cleavage by Lanthanide(III) and Cobalt(III) Derivatives. *J. Chem. Soc., Perkin Trans. 2* **1997**, 2069−2072.

(19) Kuzuya, A.; Mizoguchi, R.; Morisawa, F.; Machida, K.; Komiyama, M. Metal Ion-Induced Site-Selective RNA Hydrolysis by Use of Acridine-Bearing Oligonucleotide as Cofactor. *J. Am. Chem. Soc.* **2002**, *124*, 6887−6894.

(20) Kuzuya, A.; Machida, K.; Mizoguchi, R.; Komiyama, M. Conjugation of Various Acridines to DNA for Site-Selective RNA Scission by Lanthanide Ion. *Bioconjugate Chem.* **2002**, *13*, 365−369.

(21) Magda, D.; Miller, R. A.; Sessler, J. L.; Iverson, B. L. Site-Specific Hydrolysis of RNA by Europium(III) Texaphyrin Conjugated to a Synthetic Oligodeoxyribonucleotide. *J. Am. Chem. Soc.* **1994**, *116*, 7439−7440.

(22) Matsumura, K.; Endo, M.; Komiyama. M. Lanthanide Complex-Oligo-DNA Hybrid for Sequence-Selective Hydrolysis of RNA. *J. Chem. Soc., Chem. Commun.* **1994**, 2019−2020.

(23) Kuzuya, A.; Komiyama, M. Non-Covalent Ternary Systems (DNA-Acridine Hybride/DNA/Lanthanide(III) for Efficient and Site-Selective RNA Scission. *Chem. Commun.* **2000**, 2019−2020.

(24) Peluffo, F.; Torres, J.; Kremer, C.; Domínguez, S.; Mederos, A.; Kremer, E. Phosphodiesterolytic Activity of Samarium(III) Mixed Ligand Complexes Containing Crown Ethers and α-Amino Acids. *Inorg. Chem. Acta* **2006**, *359*, 2107−2114.

(25) Wahlroos, Ö.; Virtanen, A. L. The Precursors of 6-Methoxy-Benzoxazolinone in Maize and Wheat plants, Their Isolation and Some of Their Properties. *Acta Chem. Scand.* **1959**, *13*, 1906−1908.

(26) Streeter, C. L.; Oltjen, R. R.; Slyter, L. L.; Fishbein, W. N. Ureautilization in Wethers Receiving The Urease Inhibitor, Acetohydroxamic Acid. *J. Anim. Sci.* **1969**, *29*, 88−93.

(27) Summerskill, W. N.; Thorsell, F.; Feinberg, J.; Alderte, J. S. Effects of Urease Inhibition in Hyperammonemia: Clinical and Experimental Studies with Acetohdroxamic Acid. *Gastroenterology* **1968**, *54* (1), 20−26.

(28) Fishbein, W. N.; Daly, J. E. Urease Inhibitor for Hepatic Coma: Comparative Efficiency of Four Lower Hydroxamate Homologes in Vitro and in Vivo. *Proc. Soc. Exp. Biol. Med*. **1970**, *134*, 1083−1090.

(29) Khalifa, A. S.; Sallam, T. H.; Ghaleb, H.; Corden, B. J.; Razek, A. A.; Shoukry, M. Salicylhydroxamic Acid (SHAM): a New Oral Chelator in Transfusion Induced Hemosiderosis in Beta Thalassemia Major. *Int. J. Pediatr. Haematol./Oncol.* **1994**, *1*, 315−321.

(30) Corden, B. J. Hemosiderosis in Rodents and the Effect of Acetohydroxamic Acid on Urinary Iron Excrection. *Exp. Hematol.* **1986**, *14*, 971−974.

(31) Niemeyer, H. M.; Pesel, E., Copaja, S. V., Bravo, H. R.; Franke, S.; Francke, W. Changes in Hydroxamic Acid Levels of Wheat Plants Induced by Aphid Feeding. *Phytochemistry* **1989**, *28* (2), 447−449.

(32) Moore, E. C. The Effects of Ferrous Ion and Dithioerythritol on Inhibition by Hydroxyurea of Ribonucleotide Reductase. *Cancer Res.* **1969**, *29*, 291−295.

(33) Khalil, M. M. Complexation Equilibria and Determination of Stability Constants of Binary and Ternary Complexes with Ribonucleotides (AMP, ADP, and ATP) and Salicylhydroxamic Acid as Ligands. *J. Chem. Eng. Data* **2000**, *45*, 70−74.

(34) Khalil, M. M.; Fazary, A. E. Potentiometric Studies on Binary and Ternary Complexes of Di- and Trivalent Metal Ions Involving Some Hydroxamic Acids, Amino Acids, and Nucleic Acid Components. *Monatsh. Chem.* **2004**, *135*, 1455−1474.

(35) Khalil, M. M.; Radalla, A. M. Binary and Ternary Complexes of Inosine. *Talanta* **1998**, *46*, 53−61.

(36) Khalil, M. M.; Attia, A. E. Potentiometric Studies on the Formation Equilibria of Binary and Ternary Complexes of Some Metal Ions with Dipicolinic Acid and Amino Acids. *J. Chem. Eng. Data* **2000**, *45*, 1108−1111.

(37) Gran, G. Determination of the Equivalence Point in Potentiometric Titration. *Analyst* **1952**, *77*, 661−671.

(38) Welcher, F. J. *The Analytical Uses of Ethylenediaminetetraacetic Acid*; Von Nostrand: Princeton, NJ, 1965.

(39) Gans, P.; O'Sullivan, B. Glee, a New Computer Program for Glass Electrode Calibration. *Talanta* **2000**, *51*, 33−37.

(40) Irving, H. M.; Rossotti, H. S. Methods for Computing Successive Stability Constants from Experimental Formation Curves. *J. Chem. Soc*. **1953**, 3397−3405.

(41) Irving, H. M.; Rossotti, H. S. The Calculation of Formation Curves of Metal Complexes from *pH* Ttitration Curves in Mixed Solvent. *J. Chem. Soc*. **1954**, 2904.

(42) Douhèret, G. Liquid Junction Potentials and Effects of the Mixed Solvents (Water-dipolar a protic solvent). Application to the standardization of Glass-Calomel Electrode in Such Mixtures. *Bull. Soc. Chem. Fr*. **1968**, 3122−3131.

(43) O'Brien, E. C.; Farkas, E.; Gil, M. J.; Fitzgerald, D.; Castineras, A.; Nolan, K. B. Metal Complexes of Salicylhydroxamic Acid (H_2Sha), Anthranilic Hydroxamic Acid and Benzohydroxamic Acid. Crystal and Molecular Structure of $[Cu(phen)_2(Cl)]Cl \cdot H_2Sha$, a Model for a Peroxidase−Inhibitor Complex. *J. Inorg. Biochem.* **2000**, *79*, 47−51.

(44) Leporati, E. Acid−Base and Complex Formation Properties of *N*-Hydroxy-3-pyridinecarboxamide and *N*,2-Dihydroxybennzamide in Aqueous Solution. *Bull. Soc. Jpn*. **1991**, *64*, 3673−3676.

(45) Lee, J. D. *Concise Inorganic Chemistry*, 4th ed; Chapmann and Hall: London, U.K., 1991; Chapter 5, p 887.

(46) Stagg, W. R.; Powell, J. E. Complexes of the Trivalent Rare Earths with Isobutyrate, α-Hydroxyisobutyrate, and α,*β*,*β*′-Trihydroxyisobutyrate Ligands. *Inorg. Chem.* **1964**, *3*, 242−245.

(47) Gianguzza, A.; Milea, D.; Millero, F. J.; Sammartano, S. Hydrolysis and Chemical Speciation of Dioxouranium(IV) Ion in Aqueous Media Simulating the Major Ion Composition of Seawater. *Mar. Chem.* **2004**, *85*, 103−124.

(48) Gans, P.; Vacca, A. Miniquad−A General Computer Programme for the Computation of Formation Constants from Potentiometric Data. *Talanta* **1974**, *21*, 54−77.

(49) Ferri, D.; Luliano, M.; Manfredi, C.; Vasca, E.; Caruso, T.; Clemente, M.; Fontanella, C. Dioxouranium(VI) Oxalate Complexes. *J. Chem. Soc., Dalton Trans.* **2000**, 3460−3466.

(50) Rajan, K. S.; Martell, A. E. Equilibrium Studies of Uranyl Complexes−IV Reactions with Carboxylic Acids. *J. Inorg. Nucl. Chem*. **1967**, *29* (2), 523−529.

(51) Miyake, C.; Nürnberg, H. W. Co-ordination Compound of Actinids−I The Determination of the Stability Constants of Uranyl Complexes with Anions of Carboxylic Acids. *J. Inorg. Nucl. Chem*. **1967**, *29* (9), 2411−2429.

(52) De Stefano, C.; Gianguzza, A.; Leggio, T.; Sammartano, S. Dependance on Ionic Strength of the Hydrolysis Constants for Dioxouranium(VI) in NaCl(aq) and $NaNO_3$(aq), at pH <6 and $t = 25$ °C. *J. Chem. Eng. Data* **2002**, *47*, 533−538.

(53) Sanaie, N.; Hayres, C. A. Formation Constants and Coordination Thermodynamics for Binary and Ternary Complexes of Copper(II), L-Hydroxyproline, and an Amino Acid Enantiomer. *J. Chem. Eng. Data* **2005**, *50*, 1848−1856.

(54) Offiong, O. E. Formation Constants and Thermodynamic Parameters of α-Pyridoin Thiosemicarbazones with Divalent Metal Ions. *Trans. Met. Chem*. **1998**, *23*, 553−556.

(55) Khalil, M. M.; Taha, M. Equilibrium Studies of Binary and Ternary Complexes Involving Tricine and Some Selected α-Amino Acids. *Monatsh. Chem.* **2004**, *135*, 385−395.

(56) Fuoss, R. M.; Kraus, C. A. Properties of Electrolytic Solutions IV. The Conductance Minimum and the Formation of Triple Ions Due to the Action of Coulomb Forces. *J. Am. Chem. Soc*. **1933**, *55*, 2387−2399.

(57) Fuoss, R. M.; Kraus, C. A. Properties of Electrolytic Solutions. XV. Thermodynamic Properties of Very Weak Electrolytes. *J. Am. Chem. Soc*. **1935**, *57*, 1−4.

(58) Coetzee, J. F.; Ritchie, C. D. *Solute−Solvent Interactions*; Dekker: New York, 1969.

(59) Mui, K. K.; Mcbreyde, W. A. E.; Nieboer, E. The Stability of Some Metal Complexes in Mixed Solvent. *Can. J. Chem*. **1974**, *52*, 1821−1833.

(60) Franks, F.; Ives, D. J. G. Structural Properties of Alcohols−Water Mixtures. *Q. Rev*. **1966**, *20*, 1−44.

Received for review November 28, 2006. Accepted June 22, 2007.

JE600541A