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Interaction of Tungsten(VI) and Molybdenum(VI)
with Nitrilotriacetic Acid and Glutamic Acid
in Different Sodium Perchlorate Aqueous Solutions¹

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Abstract—Using potentiometric and spectrophotometric techniques, the complexation of tungsten(VI) with nitrilotriacetic acid (NTA) has been carried out in an aqueous solution for pH = 7.5 at 25°C and different ionic strengths ranging from 0.1 to 1.0 mol/l NaClO₄. The composition of the complex was determined by the continuous variations method. It was shown that tungsten(VI) forms a mononuclear complex with NTA of the type WO₃L³⁻ at pH = 7.5. The complexation of molybdenum(VI) with NTA and glutamic acid has been studied using potentiometric and spectrophotometric techniques. The same conditions have been used for these two complexes except that the pH was 6.0. The polarimetric technique also confirmed the complexation of molybdenum(VI) with glutamic acid. In all the aforementioned complex formation reactions, the dependence of the dissociation and stability constants on the ionic strength is described by a Debye–Hückel type equation. Finally, a comparison of the ionic strength dependence has been made.

INTRODUCTION

Desulfurovibrio gigas formate dehydrogenase is the first representative of a tungsten-containing enzyme from a mesophile that has been structurally characterized [1]. It is a heterodimer of 110 and 24 kDa subunits. The existence of a universal pterin dithiolene cofactor ligand for the molybdenum and tungsten oxotransferases supports the biological significance of the fundamental chemistry of mono- and bis(dithiolene) complexes of these elements [2]. Tungsten(VI) also forms complexes with tryptophan [3] and porphyrin [4].

Molybdenum is present in cofactors of several enzymes. Some of them are involved in important biochemical processes such as nitrogen fixation. As a result, structural, thermodynamic, and kinetic studies of the complexation between molybdate and different chelating ligands have been the subject of several investigations.

Considerable studies have been performed on the stability constants of metals with α -aminoacids and aminopolycarboxylic acids, but only a little work has been reported on the ionic strength dependence of the stability constants of tungsten(VI) and molybdenum(VI) with nitrilotriacetic acid and glutamic acid [5–7]. Marcu *et al.* [8] studied the radiochromatographic and electroradiochromatographic of sodium tungstate solutions under the action of NTA. Studies on the struc-

tural and bonding characteristics of various Mo(VI)-aminopolycarboxylic acid complexes has led to the evaluation of the stability constants from proton nuclear magnetic resonance (NMR) data [9, 10]. On the basis of these studies, Kula *et al.* [11] determined the stability constants of W(VI)-NTA complexes by potentiometric techniques. Collin *et al.* [12] reported the stability constant of this system at 25°C and at an ionic strength of 0.5 mol/l NaClO₄. Zare *et al.* [13] studied the W(VI)-NTA and Mo(VI)-NTA systems and evaluated their stability constants at 25°C and at an ionic strength of 3 mol/l NaClO₄ using the potentiometric technique. Chan *et al.* [14] proved that stable 1 : 1 molybdenum chelate is formed with NTA using the NMR technique. Raymond [15] confirmed the existence of an MoO₃ core for the complexation of Mo(VI) with the tridentate ligand NTA. Funahashi *et al.* [16] studied the reaction of molybdate(VI) with nitrilotriacetate spectrophotometrically in an aqueous solution of pH 6–8 at $I = 1.00$ M NaClO₄ and 25°C. In a potentiometric study of molybdenum(VI) chelates with glutamic acid, Rabenstein *et al.* [17] reported the stability constant at 25°C and at an ionic strength of 0.2 mol/l KNO₃. Raymond *et al.* [15] synthesized and proved the composition of the complex using a spectroscopic technique. Gharib *et al.* [18] have proved the composition of a glutamic acid complex and reported its stability constant using the polarimetric and spectrophotometric techniques.

¹ The text was submitted by the authors in English.

Table 1. Dissociation constants K_3 , K_2 , and K_1 of NTA at different ionic strengths, I , of NaClO₄

I , mol/l	$\log K_3$	$\log K_2$	$\log K_1$	Experimental Conditions	Ref
0.1	1.98 ± 0.02	2.92 ± 0.05	10.00 ± 0.02		This work
0.3	1.83 ± 0.04	2.84 ± 0.03	9.85 ± 0.01		This work
0.5	1.76 ± 0.01	2.79 ± 0.02	9.70 ± 0.03		This work
0.7	1.63 ± 0.02	2.61 ± 0.01	9.55 ± 0.02		This work
1.0	1.55 ± 0.02	2.59 ± 0.04	9.25 ± 0.02		This work
			9.81 ± 0.10	$I = 0.15 \text{ M}, T = 25^\circ\text{C}$	11
	2.05 ± 0.05	2.63 ± 0.02	9.17 ± 0.04	$I = 3 \text{ M NaClO}_4, T = 25^\circ\text{C}$	13
	1.65	2.94	10.33	$T = 20^\circ\text{C}$	34

The present work deals with the complexation of tungsten(VI) and molybdenum(VI) with NTA and molybdenum(VI) with glutamic acid in an ionic strength range of 0.1–1.0 mol/l sodium perchlorate at 25°C. A simple Debye–Hückel type equation was established for the dependence of the formation constants on the ionic strength. This equation makes it possible to estimate a stability constant at a fixed ionic strength when its value is known in another media in the range 0.1 mol/l I < 1.0 mol/l.

EXPERIMENTAL

Reagents. Sodium perchlorate, perchloric acid, sodium hydroxide, sodium tungstate, sodium molybdate, nitrilotriacetic acid, and glutamic acid were obtained from Merck as analytical reagent grade materials and were used without further purification. Dilute perchloric acid solution was standardized against KHCO₃. In all the experiments, double-distilled water with a specific conductance equal to $1.3 \pm 0.1 \mu\Omega^{-1} \text{ cm}^{-1}$ was used.

Measurements. A Horiba pH meter (D-14) was used for the pH measurements. The pH meter has a sensitivity of 0.01. The hydrogen ion concentration was measured with a Horiba combination electrode (model S8720). A 0.01 mol/l perchloric acid solution containing 0.09 mol/l sodium perchlorate (for adjusting the ionic strength to 0.1 mol/l) was employed as a standard solution of hydrogen ion concentration. The same procedure was performed for the other ionic strengths [13]. The calibration was done for the whole pH ($\text{pH} = -\log[\text{H}^+]$) range used. Spectrophotometric measurements were performed on a UV–Vis Shimadzu 2101 spectrophotometer with an Acer Mate 486 SX/250 computer using thermostated, matched 10-mm quartz cells. Polarimetric measurements were performed with an Atago model Polax-D polarimeter equipped with a

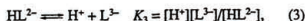
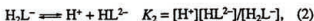
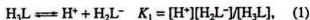
sodium lamp. A water-jacketed cell of 200 mm length and 20 cm³ total volume was used.

For each experiment, two solutions of metal + ligand were prepared with the same concentration, but the ionic strength of the first was maintained with sodium perchlorate and that of the second with sodium hydroxide or perchloric acid. The pH of the first solution was adjusted with the second one. The first solution consisted of metal + ligand + NaOH for increasing the pH, while, for decreasing the pH, the second one consisted of metal + ligand + HClO₄. The absorbance of the first solution was measured after adjusting the pH.

In all cases, the procedure was repeated at least three times, and the resulting average values and corresponding standard deviations are shown in the text and tables.

RESULTS AND DISCUSSION

The dissociation equilibria of nitrilotriacetic acid and glutamic acid have been studied in different kinds of background electrolytes, but there are no reports about the ionic strength dependence of the dissociation constants of NTA and glutamic acid [5–7]. The following equilibria were studied:



where L^{3-} represents the fully dissociated ligand anion. The dissociation constants K_1 , K_2 , and K_3 have been determined using potentiometric techniques and calculated using the Solver and Microsoft Excel 2000 powerful optimization package to perform nonlinear least-squares curve fitting [19, 20]. For the glutamic acid,

Table 2. Dissociation constants K_2 and K_1 of *L*-glutamic acid at different ionic strengths, I , of NaClO₄

I , mol/l	$\log K_2$	$\log K_1$	Experimental Conditions	Ref
0.1	4.24 ± 0.05	9.64 ± 0.05		This work
0.3	4.03 ± 0.05	9.27 ± 0.05		This work
0.5	3.76 ± 0.05	9.19 ± 0.05		This work
0.7	3.74 ± 0.05	9.04 ± 0.05		This work
1.0	3.60 ± 0.05	8.98 ± 0.05		This work
	4.15	9.49	$I = 0.15 \text{ M NaClO}_4, T = 25^\circ\text{C}$	18
	4.21	9.54	$I = 0.1 \text{ M NaNO}_3, T = 25^\circ\text{C}$	35
	4.05	9.46	$I = 0.1 \text{ M NaClO}_4, T = 30^\circ\text{C}$	36
	4.15	9.61	$I = 0.1 \text{ M KNO}_3, T = 25^\circ\text{C}$	37
	3.71	9.63	$I = 0.1 \text{ M NaClO}_4, T = 30^\circ\text{C}$	38

only K_1 and K_2 have been determined. These values are listed in Tables 1 and 2, together with the values reported in the literature, which are in good agreement with those reported before.

Complexation of Tungsten(VI) with NTA

Using the continuous variations method, we determined the absorbances of solutions of W(VI) and NTA of total concentration 0.006 mol/l in the UV range 260–265 nm at a constant pH of 7.5. When solutions of tungstates are made weakly acid, polymeric anions are formed, but, from more strongly acid solutions, substances often called tungstic acid are obtained [21]. The behavior of the tungstate systems is similar to that of the molybdate systems. Again, the degree of aggregation in the solution increases as the pH is lowered, and numerous tungstates have been crystallized from the solutions at different pH [21].

It is now certain that the WO_4^{2-} ion is tetrahedral in aqueous solution [11]. In the usual potentiometric method for evaluating metal–ligand stability constants, the competition between the metal ion and hydrogen ion for the ligand is studied and the pH region of interest is from about 1 to 5 [11]. In the W(VI) systems, however, the complication of metal polymerization is introduced in acidic solutions. Because the polymerization equilibria are not well understood, this pH region is not useful for stability determinations. In more alkaline solutions, on the other hand, a pH-dependent process involving the competition between tungstate formation and metal–ligand complexation, i.e., a competition between OH^- and ligand for the metal ion, can be utilized. This process was determined from the NMR studies to be important from about pH 6 to 9 and can be represented by [11]:



where M represents W and L represents the aminopolycarboxylic acid ligand. In the pH region above 6, no evidence was found for any Mo species containing fewer than three oxygen atoms, e.g., MoO_2^{2+} , as has been proposed for other systems [22]. The molybdenum coordinating species in all the aminopolycarboxylic acid systems above pH 6 is MoO_3 , and, by analogy, we have assumed that the corresponding coordinating unit in the tungsten systems is WO_3 .

Using the potentiometric technique, results obtained for simple 1 : 1 metal–ligand chelates are more precise than for higher complexes, such as those formed with EDTA. In fact, for the multicomplex systems, the NMR data are probably more reliable than the potentiometric data as indicated by the range of calculated values for the EDTA systems [11]. A comparison of the formation constants shows that, within the experimental error, there is essentially no difference between the stabilities of the corresponding Mo(VI) and W(VI) chelates. This does not seem too surprising in view of the similarities of the two ions and their nearly equal ionic radii [11]. However, NMR studies show that tungsten chelates are significantly more labile with respect to individual metal–ligand bonds than the molybdenum chelates are. The slowness with which the pH equilibrium was attained in the NTA chelates may be explained by the higher negative charge of this complex compared to the other complexes and the subsequently slower reaction with OH^- :



For the NTA ligand system, NMR studies also indicate that only one metal–ligand species exists above pH 6,

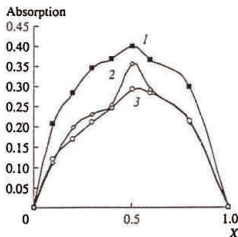


Fig. 1. Continuous variations plots of the absorbances of (1) $\text{MoO}_3\text{NTA}^{3-}$, (2) $\text{WO}_3\text{NTA}^{3-}$, and (3) $\text{MoO}_3\text{Glu}^{2-}$, versus the mole fractions of W(VI) and Mo(VI) , X , at 25°C , an ionic strength of $0.1 \text{ mol/l NaClO}_4$ and 265 nm .

Mo_3L^{3-} [11]. Thus, W(VI) will bond with this tridentate ligand as a 1 : 1 complex:



with the stability constant K_s as

$$K_s = [\text{WO}_3\text{NTA}^{3-}] / [\text{WO}_4^{2-}][\text{NTA}^{3-}][\text{H}^+]^2. \quad (7)$$

The values of $\log K_s$ at different ionic strengths,

Table 3. Average values of $\log K_s$ at pH 7.5 and different ionic strengths for the complexation of tungsten(VI) with NTA, $T = 25^\circ\text{C}$

$I, \text{mol/l}$	$\log K_s$	Experimental conditions	Ref.
0.1	19.00 ± 0.30		This work
0.3	19.37 ± 0.10		This work
0.5	19.65 ± 0.25		This work
0.7	20.10 ± 0.10		This work
1.0	19.40 ± 0.20		This work
	18.86 ± 0.05	$I = 0.15 \text{ M}$, $T = 25^\circ\text{C}$	11
	19.10 ± 0.2	$I = 1.0\text{--}2.5 \text{ M}$, $T = 35^\circ\text{C}$	11
	17.75	$I = 0.5 \text{ M NaClO}_4$, $T = 25^\circ\text{C}$	12
	19.03 ± 0.15	$I = 3 \text{ M NaClO}_4$, $T = 25^\circ\text{C}$	13

together with the values of the literature, are shown in Table 3. These values have been calculated similar to the molybdenum(VI) complex. Calculations for the molybdenum complex are described in the next section.

Complexation of Molybdenum(VI) with Nitriiotriacetic Acid and Glutamic Acid

1. Polarimetric studies. The effect of pH on the optical activity of glutamic acid and its complex with molybdenum(VI) shows the difference of optical rotation for the ligand and the complex. This difference reaches a maximum in the pH range 5.0–6.0, which means that we have the largest amount of complex formation in this pH range.

2. Spectrophotometric studies. Using the continuous variations method, we determined the absorbances of solutions of Mo(VI) and NTA of total concentration 0.006 mol/l in the UV range $260\text{--}265 \text{ nm}$ at a constant pH of 6. The total concentration for glutamic acid was 0.02 mol/l . Different sets of species have been proposed in order to assign the equilibria in molybdate solutions. The equilibria of molybdenum(VI) in acidified molybdate solutions are complex since various polynuclear species are formed [16]. At higher pH, the complex dissociates as a result of the competitive formation of MoO_4^{2-} [21]. The observed absorbances were corrected from Eq. (8) and plotted in Fig. 1:

$$A_c = A_{\text{obs}} - \epsilon_0[\text{metal}], \quad (8)$$

where A_c , A_{obs} , and ϵ_0 are the absorbance of the complex, the observed absorbance, and the molar absorptivity of the metal, respectively. The ϵ_0 values were calculated at the mole fraction of the metal equal to 1 and are shown in Table 4 for tungsten(VI) and Table 5 for molybdenum(VI). In Fig. 1, a maximum at a mole fraction of the metal equal to 0.5 was obtained, indicating a 1 : 1 complex. The molar absorptivity of the complex, ϵ_1 was calculated from the linear part of the aforementioned plot at a low mole fraction of the metal, where, essentially, all the metal ions were in the form of a com-

Table 4. Molar absorptivities of W(VI) , ϵ_0 , and $\text{WO}_3\text{NTA}^{3-}$, ϵ_1 , at pH 7.5, different wavelengths, and various ionic strengths, I , of NaClO_4

$I, \text{mol/l}$	ϵ_0		ϵ_1	
	260 nm	265 nm	260 nm	265 nm
0.1	660.8	486.6	1116.6	1116.6
0.3	527.5	365.0	1221.6	1268.3
0.5	505.8	345.0	1533.3	1583.3
0.7	478.3	320.0	716.6	750.0
1.0	522.5	355.0	2316.6	2350.0

Table 5. Molar absorptivities of Mo(VI), ϵ_0 , and MoO₃NTA³⁻, ϵ_1 , for the Mo(VI) + NTA system at pH 6, different wavelengths and various ionic strengths, I , NaClO₄

I , mol/l	ϵ_0		ϵ_1	
	260 nm	265 nm	260 nm	265 nm
0.1	388.3	205.8	1928.3	1726.6
0.3	346.7	166.7	1166.6	983.3
0.5	380.0	199.2	3066.6	2916.6
0.7	386.7	211.7	2346.6	2188.3
1.0	402.5	223.3	1830.0	1710.0

Table 7. Average values of $\log K_s$ at pH 6 and different ionic strengths, for the complexation of molybdenum(VI) with NTA, $T = 25^\circ\text{C}$

I , mol/l	$\log K_s$	Experimental conditions	Ref.
0.1	18.72 ± 0.20	$I = 3 \text{ M NaClO}_4$, $T = 25^\circ\text{C}$	This work
0.3	19.36 ± 0.10		This work
0.5	18.08 ± 0.30		This work
0.7	18.31 ± 0.20		This work
1.0	17.97 ± 0.10		This work
	18.60 ± 0.20	$I = 3 \text{ M NaClO}_4$, $T = 25^\circ\text{C}$	13
	18.09 ± 0.04	$I = 1 \text{ M NaClO}_4$, $T = 25^\circ\text{C}$	16

Table 6. Molar absorptivities of MoO₃Glu²⁻, ϵ_1 , for the Mo(VI) + L-glutamic acid system at pH 6, different wavelengths and various ionic strengths, I , NaClO₄

I , mol/l	ϵ_1	
	260 nm	265 nm
0.1	273.5	98.8
0.3	285.1	109.5
0.5	268.1	94.4
0.7	269.9	97.4
1.0	265.1	98.8

Table 8. Average values of $\log K_s$ at pH 6 and different ionic strengths, for the complexation of molybdenum(VI) with L-glutamic acid, $T = 25^\circ\text{C}$

I , mol/l	$\log K_s$	Experimental conditions	Ref.
0.1	17.54 ± 0.30	$I = 0.2 \text{ M NaClO}_4$, $T = 25^\circ\text{C}$	This work
0.3	16.94 ± 0.40		This work
0.5	16.93 ± 0.50		This work
0.7	16.84 ± 0.35		This work
1.0	16.76 ± 0.40		This work
	16.78	$I = 0.2 \text{ M NaClO}_4$, $T = 25^\circ\text{C}$	[17]
	16.73	$I = 0.15 \text{ M NaClO}_4$, $T = 25^\circ\text{C}$	[18]

plex and are listed in Tables 4–6. At the maximum point of the plot, the concentration of the complex is

$$[C] = A_0/\epsilon_1 \quad (9)$$

The molybdate anion MoO₄²⁻ maintains a tetrahedral configuration in neutral and alkaline solutions [21]. Complexes of molybdate with chelating ligands have an octahedral configuration. The complex formation, therefore, will have to occur by an addition of the ligand to the molybdate ion, thereby increasing its coordination number from 4 to 6. It has been postulated that the monoprotonated species HMoO₄⁻ exists in the form of an octahedral hydrate species in solution. From consideration of the thermodynamic parameters for protonation of molybdate, Cruywagen *et al.* [23] have suggested that it is the second protonation constant that is anomalous and that the change in coordination number occurs with addition of the second proton. Whichever of these viewpoints is correct, the diprotonated species H₂MoO₄ should be octahedral.

Although Mo(VI) forms complexes with the same chelating agents as most other metal ions, its chemistry is differentiated from other transition ions by its strong association with oxygen. Thus, in most Mo(VI) complexes, MoO₂²⁺ or MoO₃ is the central coordinating unit, which, with octahedral geometry, severely limits the number of coordination sites available to the ligands. An interesting consequence of this behavior is illustrated by the complex formed between Mo(VI) and EDTA in which two Mo ions can coordinate with EDTA [24]. This is in distinct contrast to the normal metal–EDTA chelates in which only one metal ion coordinates with each ligand. Kula [10] concluded that, for EDTA (between pH 9 and 5), two complexes with a 1 : 1 and 2 : 1 metal–ligand ratio are formed. The structures that were proposed for the Mo₂–EDTA chelate led to speculation concerning the possibility of forming 1 : 1 Mo(VI) chelates with methyliminodiacetic acid (MIDA) and nitrilotriacetic acid (NTA), and proton NMR studies of these chelates confirmed that the predominate EDTA chelate does indeed contain two Mo

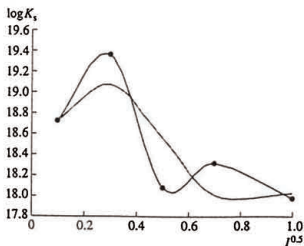


Fig. 2. Plot of $\log K_s$ for $\text{MoO}_3\text{NTA}^{3-}$ versus the square root of the ionic strength.

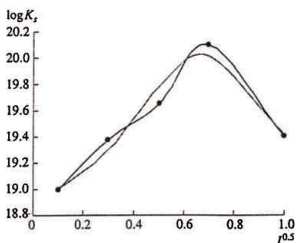


Fig. 3. Plot of $\log K_s$ for $\text{WO}_3\text{NTA}^{3-}$ versus the square root of the ionic strength.

ions and that stable 1 : 1 molybdenum chelates are formed with MIDA and NTA [9, 24]. Thus, Mo(VI) will bond with this tridentate ligand (NTA) as a 1 : 1 complex:



with the stability constant K_s as

$$K_s = [\text{MoO}_3\text{NTA}^{3-}]/[\text{MoO}_4^{2-}][\text{NTA}^{3-}][\text{H}^+]^2. \quad (11)$$

For the glutamic acid, the species is $\text{MoO}_3\text{Glu}^{2-}$ with

Table 9. Parameters for the dependence on the ionic strength of dissociation and stability constants for the W(VI) + NTA and Mo(VI) + NTA systems at 25°C

Species	C	D	E	Z*
K_3	2.933	-4.807	2.161	6
K_2	4.860	-9.257	4.673	4
K_1	-0.292	0.280	-0.578	2
$\text{WO}_3\text{NTA}^{3-}$	-8.294	28.021	-18.886	6
$\text{MoO}_3\text{NTA}^{3-}$	33.961	-67.906	35.411	6

Table 10. Parameters for the dependence on the ionic strength of dissociation and stability constants for the Mo(VI) + L-glutamic acid system at 25°C

Species	C	D	E	Z*
K_2	-0.782	-0.453	0.702	2
K_1	-3.346	5.097	-2.241	4
$\text{MoO}_3\text{Glu}^{2-}$	-10.792	19.924	-9.911	6

the stability constant as

$$K_s = [\text{MoO}_3\text{Glu}^{2-}]/[\text{MoO}_4^{2-}][\text{Glu}^{2-}][\text{H}^+]^2. \quad (12)$$

The following equations are valid for the total concentration of the metal (c_M) and the total concentration of the ligand (c_L) at the maximum point on the plot of Fig. 1:

$$c_M = [\text{metal}] + [\text{C}], \quad (13)$$

$$c_L = [\text{L}] + [\text{C}]. \quad (14)$$

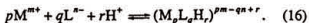
By substituting Eqs. (9), (13), and (14) in Eqs. (7), (11), and (12), we can calculate the values of K_s . The values of $\log K_s$ at different ionic strengths, together with the values of the literature, are shown in Tables 3, 7, and 8.

Ionic Strength Dependence of Dissociation and Stability Constants

The dependence of the dissociation and stability constants on the ionic strength can be described according to the previous works [5–7, 20, 25–33]:

$$\log K_s(I) = \log K_s(I_1) - AZ^* \left(\frac{I^{0.5}}{1 + BI^{0.5}} - \frac{I_1^{0.5}}{1 + BI_1^{0.5}} \right) + C(I - I_1) + D(I^{1.5} - I_1^{1.5}) + E(I^2 - I_1^2) \quad (15)$$

where I and I_1 are the actual and reference ionic strengths, respectively, and, according to Eq. (16),



$Z^* = pm^2 + qn^2 + r - (pm + qn + r)^2$, where m and n are the charges on the metal ion and the ligand, respec-

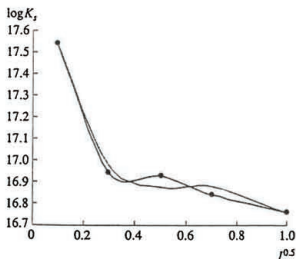


Fig. 4. Plot of $\log K_s$ for $\text{MoO}_3\text{Glu}^{2-}$ versus the square root of the ionic strength.

tively. Considering $A = 0.5115$ and $B = 1.489$, Eq. (15) can be simplified:

$$\log K_s(I) = \log K_s(I_1) - Z^* \left(\frac{I^{0.5}}{2 + 3I^{0.5}} - \frac{I_1^{0.5}}{2 + 3I_1^{0.5}} \right) + C(I - I_1) + D(I^{1.5} - I_1^{1.5}) + E(I^2 - I_1^2), \quad (17)$$

where C , D , and E are empirical coefficients, and their values were obtained by minimizing the error squares sum, U , and the Gauss-Newton nonlinear least-squares method in a suitable computer program:

$$U = \sum_i (a_i - b_i)^2 \quad (i = 1, 2, 3, \dots), \quad (18)$$

where a is a quasi-experimental quantity and b_i is a calculated one. The values of C , D , and E are shown in Tables 9 and 10.

CONCLUSIONS

Sodium molybdate has three sites to form its complexes in aqueous solution. The same situation exists for sodium tungstate. On the other hand, glutamic acid and nitrilotriacetic acid are potentially tridentate ligands with three donor sites. Thus, Mo(VI) and W(VI) will bond with these tridentate ligands, and it seems unlikely that the complexes of 1 : 2 and 1 : 3 stoichiometry exist in the pH which has been used.

Tables 3 and 7 show that, in both the $\text{W(VI)} + \text{NTA}$ and $\text{Mo(VI)} + \text{NTA}$ systems, a maximum at $I = 0.7$ is obvious, but another maximum can be seen at $I = 0.3$ for the latter. In the $\text{Mo(VI)} + \text{glutamic acid}$ system, no maximum can be detected. It seems that, at the minimum points, the attraction and repulsive forces between

the ions are equal. The attraction forces are dominant before the minimum points, and the systems are stabilized. The repulsive forces are more important after the minimum points; therefore, the stability of the systems is decreased.

According to the obtained values of C , D , and E and Eq. (17), it is possible to estimate a stability constant at a fixed ionic strength when its value is known in another ionic media in the desired range. In Figs. 2–4, the dot and continuous curves are the values of the experimental and calculated stability constants, respectively.

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