International Journal of Nanomaterials and Chemistry

http://dx.doi.org/10.12785/ijnc/010201

Stability Studies of Selected Metal Ions Chelates with 2-(4-amino-1,5-dimethyl-2-phenyl-1,2-dihydro-pyrazol-3-ylideneamino) Phenol

Gamal A. H. Gouda [1,*], *Gomaa A. M. Ali* [1, 2] and *Tarek A. Seaf Elnasr* [1]

[1] *Chemistry Department, Faculty of Science, Al–Azhar University, Assiut, 71524, Egypt.*
2 Faculty of Industrial Sciences & Technology, University Malaysia Pahang, 26300 Kuantan, Malaysia.

Received: 12 Dec. 2014, Revised: 2 Jan. 2015, Accepted: 3 Jan. 2015.

Abstract: Acid dissociation and complex stability constants of 2-(4-Amino-1,5-dimethyl-2-phenyl-1,2-dihydro-pyrazol-3-ylideneamino) phenol (H_2A) with some metal ions chelates have been determined potentiometrically using 0.1 M KCl and 20 % (v/v) ethanol-water mixture. The stability constants of the formed complexes were found to be in the order $Re^{5+} > Ce^{3+} > Pd^{2+} > Th^{4+} > Hg^{2+} > Co^{2+}$. The dissociation constants (pK_a) of H_2A and the stability constants ($\log K_i$) of their complexes were determined at 298, 308 and 318 K. The corresponding thermodynamic parameters (ΔG, ΔH and ΔS) were calculated and discussed. The results indicated that, the dissociation process is non-spontaneous, endothermic and entropically unfavorable. The formation of the metal complexes has been found to be endothermic and entropically favorable.

Keywords: Schiff-base derivative, dissociation and stability constants, thermodynamics.

1 Introduction

Schiff-base coordination complexes have attracted a great attention over the last decades. That is due to their facile syntheses, their wide application, the accessibility of diverse structural modifications and their biological modeling applications, catalysis, design of molecular ferromagnets, and materials chemistry [1, 2]. It is well known that N-atoms play a key role in the coordination of metals at the active sites of numerous metallobiomolecules [3]. The Schiff-bases of 4-aminoantipyrine and their coordination complexes have been extensively investigated because of their biological, clinical, pharmacological, analytical and material applications [4, 5]. In addition, some drugs showed increased activity when administered as metal chelate rather than as free organic compounds [6].

Recently, metal complexes of Schiff base containing pyrazol ligands have been widely studied in the solid state by spectral characterization [7, 8]. However to best of our knowledge, no detailed studies have been carried out on the chelating tendencies of 2-(4-Amino-1,5-dimethyl-2-phenyl-1,2-dihydro-pyrazol-3-ylideneamino) phenol with bi, tri, tetra and pentavalent transition metal ions. The objective of this study are the potentiometric determination of the dissociation constants of 2-(4-Amino-1,5-dimethyl-2-phenyl-1,2-dihydro-pyrazol-3-ylidene- amino) phenol (H_2A) and the stability constant of their complexes with Co^{2+}, Pd^{2+}, Hg^{2+}, Re^{5+}, Th^{4+} and Ce^{3+} at different temperatures. Furthermore, the corresponding thermodynamic functions of complexation are evaluated and discussed.

2 Experimental

2-(4-Amino-1,5-dimethyl-2-phenyl-1,2-dihydro-pyrazol-3-ylideneamino) phenol (Figure 1) were prepared by refluxing a mixture of hot solution of 4-aminoantipyrine (10^{-3} M) and 2-hydroxyaniline (10^{-3} M) in 50 ml ethanol for 2 h [9]. The formed solid product was separated by filtration, purified by crystallization from ethanol, washed with diethyl ether, and dried in a vacuum over anhydrous calcium chloride. Then the ligand solution (1.5×10^{-3} M) was prepared by dissolving an accurately weighted amount of the solid in ethanol (Analar).

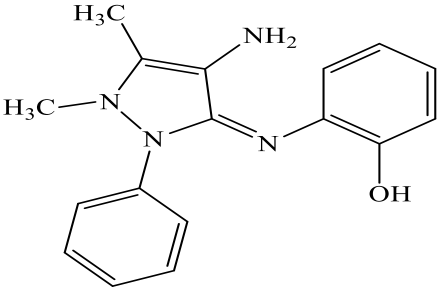

Fig. 1: Structure of 2-(4-Amino-1,5-dimethyl-2-phenyl-1,2-dihydro-pyrazol-3-ylideneamino) phenol.

Metal ion solutions of Co^{2+}, Pd^{2+}, Hg^{2+}, Th^{4+} and Ce^{3+} were prepared from Analar metal salt samples (BDH) in

---

[*] Corresponding author e-mail: *ggouda73@mail.ru*

bidistilled water and standardized with EDTA [10]. Metal ion solution of Re^{5+}, (K_2ReOCl_5) was prepared as reported elsewhere [11]. The concentration of metal ion in the solution was determined spectro-photometrically at wavelength (λ) of 480 nm. A carbonate-free sodium hydroxide solution in 20 % ethanol-water mixture was used as a titrant and standardized against oxalic acid (Analar). The calculation methods were as reported in references [12]. The following mixtures were prepared and titrated potentiometrically at three different temperatures against standard 0.2 M NaOH in a 20 % ethanol-water mixture containing 0.1 M KCl. For each solution mixture the volume was made up to 50 ml with 20 % ethanol-water mixture before the titration. The following titrations were carried out: (i) 5 ml of hydrochloric acid (HCl) (0.1 M) + 5 ml of potassium chloride (KCl) (1 M), (ii) Solution (i) + 10 ml of the ligand under investigation (1.5×10^{-3} M), and (iii) Solution (ii) + 10 ml (1.5×10^{-4} M) of metal ion.

The temperature was controlled to within ± 0.05 K by circulating thermostated water (Neslab 2 RTE 220) through the outer jacket of the vessel. The pH measurements were performed with a Metrohm 836 Titrando (KF & Potentiometric Titrator) equipped with a combined porolyte electrode. The pH-meter readings in the non-aqueous medium were corrected [13]. The electrode system was calibrated according to the method of Irving et al. [14]. The proton-ligand equilibrium constant for the ligand under experimental conditions were determined by Calvin-Bjerrum pH-titration [15], which is modified by Irving and Rossotti [16]. All titrations have been carried out between pH 2.0 and 12.0 and under nitrogen atmosphere.

3 Results and Discussion

3.1 Determination of Stability Constant of the Binary Complexes

The average number of the protons associated with H_2A at different pH values, $\bar{n}_A$ was calculated from the titration curves of the acid in the absence and presence of H_2A. The proton-ligand formation $(\bar{n}_A)$ was calculated using Irving and Rossotti [16] expression (Eq. (1)):

$$\bar{n}_A = \gamma + \frac{(V_1 - V_2)(N^o + E^o)}{(V_T + V_1)\, T_L^o} \qquad (1)$$

where, V_1 and V_2 are the volumes of alkali (ml) required to reach the same pH in acid and ligand titration curves, respectively, T_L^o is the total ligand concentration in 50 ml solution, γ is the total number of protons free attached to the ligand molecule, N^o is the normality of the alkali, E^o is the initial concentration of free acid and V_T is the total volume (ml) of the titration solution. The average number of metal ion associated with the ligand $(\bar{n})$ at different pH values was calculated from the metal ion and ligand titration curves using the following equations (2) and (3) [16]:

$$\bar{n} = \frac{(V_3 - V_2)(N^o + E^o)}{(V_T + V_2)\, \bar{n}_A\, T_M^o} \qquad (2)$$

$$pL = \log_{10}\left[\frac{\sum_{n=0}^{n=i} \beta_n^H \left(\frac{1}{\text{anti}\log pH}\right)^n}{T_L^o - \bar{n}\, T_M^o} \cdot \frac{V_T + V_3}{V_T}\right] \qquad (3)$$

where, V_3 is the volume (ml) of metal ions present in the solution, T_M^o denotes the total concentration of metal present in the solution and β_n^H is the overall proton ligand stability constant. The pH-metric titrations were carried out in a restricted range of pH values from 2.9 to 10.6, since the solutions became cloudy at higher pH values, probably due to the formation of sparingly soluble polymeric complexes or hydroxo-species [17]. The precipitates persist up to pH = 11. This phenomenon was observed with Re^{5+} and Th^{4+} complexes [18]. Dilute solutions of these ions (1.5×10^{-4} M) have been used to avoid precipitation and formation of polynuclear or hydroxo-complexes. Also, the calculations were restricted to data obtained before precipitation.

The ligand undertaken to present study contains a hydroxyl group, so it can donate a proton and the remaining anion can act as a chelating agent. When any ligand is mixed with metal ion in a solution, there is a competition between the metal ion and the protons for capturing the ligand. It is obvious that there would be a change in the pH of the solution in the formation of complex. Subsequently, the complexation reaction can be studied potentiometrically. The difference in the volume of NaOH utilized for the ligand and metal ions titration was a measure of extent of complexation. Different trends in the titration curves obtained by plotting the observed pH against the volume of alkali, as shown in fig. 2.

As the titration curve of the ligand in Fig. 2(b), it could be seen that there are two end–points at 1.66 ml and 2.63 ml, which indicate that the Schiff base (H_2A) has two protonation constants. The values of pK_{OH} and pK_{NH_2} are found to be 6.91 and 4.75, respectively at 298 K. The higher value of pK_{OH} is due to the protonation of phenolic oxygen while the lower pK_{NH_2} value is due to the amino group protonation. The metal-ligand complex stability constants were obtained by analysis of formation curves (Fig. 3) by plotting $\bar{n}$ versus p[H_2A]. When the maximum value of $\bar{n}$ obtained was more than 2, the ligand is dibasic in nature.

The dissociation constants (pk_a) for H_2A as well as the stability constants ($\log K_i$) of its complexes with selected metal ions undertaken to present study have been evaluated at 298, 308 and 318 K, and listed in Table 1.

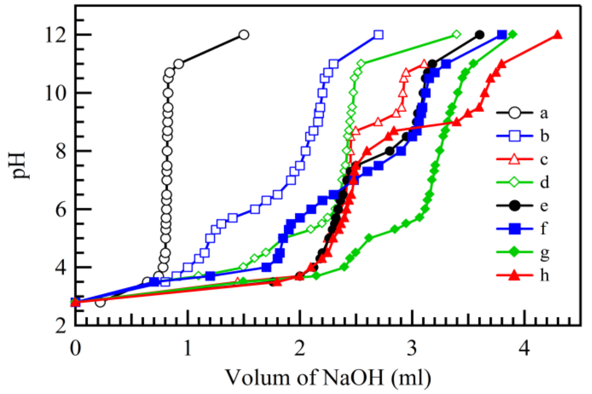

Fig. 2: Titration curves for (a) HCl, (b) HCl with H_2A, HCl with H_2A and (c) Hg^{2+}, (d) Co^{2+}, (e) Th^{4+}, (f) Pd^{2+}, (g) Re^{5+} and (h) Ce^{3+} at 298 K in at 0.1 M KCl (20 % ethanol-water mixture).

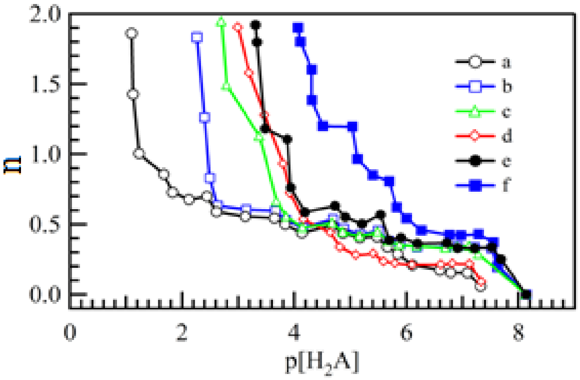

Fig. 3: Formation curves of metal ions complexes with H_2A, (a) Co^{2+}-H_2A (b) Hg^{2+}-H_2A, (c) Th^{4+}-H_2A, (d) Pd^{2+}-H_2A, (e) Ce^{3+}-H_2A and (f) Re^{5+}-H_2A systems at 298 K.

Table 1: Proton-ligand and metal-ligand stability of H_2A in 20 % ethanol-water mixture at various temperatures.

T (K)	pK_1	pK_2	$\log K_i$	Co^{2+}	Pd^{2+}	Hg^{2+}	Re^{5+}	Th^{4+}	Ce^{3+}
298	6.91	4.75	$\log K_1$	3.85	4.45	3.96	6.21	4.01	5.21
			$\log K_2$	1.12	3.29	2.34	4.31	2.92	3.39
308	6.80	4.64	$\log K_1$	3.74	4.36	3.88	6.12	3.94	5.12
			$\log K_2$	1.04	3.24	2.27	4.25	2.90	3.30
318	6.72	4.53	$\log K_1$	3.03	4.27	3.78	6.03	3.86	5.03
			$\log K_2$	0.94	3.20	2.19	4.19	2.85	3.23

The proton-ligand stability constant (pK_a) values decrease with increase in temperature, i.e. the acidity of the ligands increases [19]. This suggests that the liberation of protons becomes easier at higher temperature. The following general remarks can be pointed out: (i) The maximum value of $\bar{n}$ was around 2 indicating the formation of 1:1 and 1:2 (metal:ligand) complexes only [20]. (ii) The metal ion solution used in the present study was very dilute (1.5×10^{-4} M); hence there was no possibility of formation of polynuclear complexes [21]. (iii) The metal ions titration curves were displaced to the right hand side of the ligand titration curves along the volume axis, indicating proton release upon complex formation of the metal ion with the ligand. The large decrease in pH for the metal titration curves relative to ligand titration curves points to the formation of strong metal complexes [22].

For the same ligand at constant temperature, the stability of the chelates increases in the order $Re^{5+} > Ce^{3+} > Pd^{2+} > Th^{4+} > Hg^{2+} > Co^{2+}$. This order largely reflects the changes in the heat of complex formation across the series from a combination of the influence of both the polarizing abilities of the metal ion and the crystal-field stabilization energies [23].

The concentration distribution diagrams at 308 K of the protonated forms of the ligand (H_2A) and for the Ce^{3+} complexes as an representative are shown in figures 4 and 5, respectively. The stoichiometric protonation constants of the investigated H_2A ligand were determined in 20 % ethanol-water mixture at 308 K. The H_2A ligand studied here have two protonation constants. In acidic solution (pH < 3.2), H_2A initially exists in the fully protonated form as H_2A^{2+}. On raising the pH, the species (H_2A^{2+}) lose one of its protons forming (HA^+), which is the predominant species in pH range 4.5-6.8 with maximum concentration 86 %. The second proton is deprotonated forming the species (A) with maximum concentration percentage of 96 % at pH > 6.8.

The Ce^{3+}-(H_2A^{2+}) system starts to form at pH value 2.8, reaching a maximum concentration (80.25 %) in pH range 3.3-5.1. On the other hand, Ce^{3+}-$(A)_2$ system concentration was found to increase with increasing the pH and becomes predominant (99 %) at pH = 7.2.

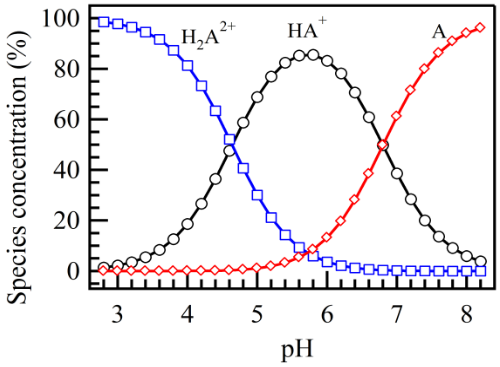

Fig.4: Species distribution diagram for H_2A (1.5 mM of H_2A) as a function of pH at 308 K.

3.2 Determination of the Thermodynamic Parameters

The thermodynamic parameters such as Gibb's free energy change (ΔG) for formation of complexes were calculated using Eq. (4). In addition, enthalpy (ΔH) was obtained from the slope of the plot log K_i against 1/T (Fig. 6) using the graphical representation of Van't Hoff equations (eq. (5)). From the ΔG and ΔH values one can obtain the entropy change (ΔS) using the eq. (6).

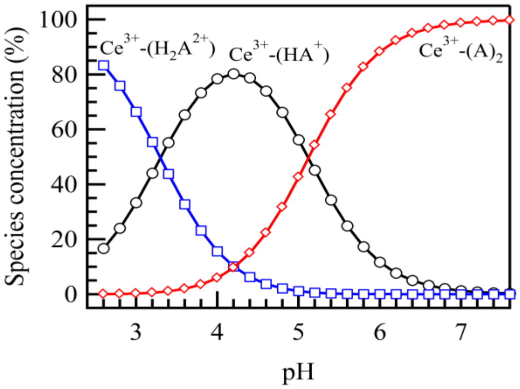

Fig. 5: Species distribution diagram for the systems Ce^{3+}–H_2A (1.5 mM of Ce^{3+} and H_2A) as a function of pH at 308 K.

$$\Delta G = -2.303\ RT \log K_i = \Delta H - T\Delta S \quad (4)$$

$$\log K_i = \left(\frac{-\Delta H}{2.303R}\right)\left(\frac{1}{T}\right) + \left(\frac{\Delta S}{2.303R}\right) \quad (5)$$

$$\Delta S = (\Delta H - \Delta G) / T \quad (6)$$

Where, R is the ideal gas constant (8.314 J/K mol), T is the absolute temperature in Kelvin (K) and K_i is the stability constant of the complex.

All the thermodynamic parameters of the stepwise stability constants of complexes are given in Table 2. The negative ΔG values indicate that both dissociation of the ligand and the complexation process are spontaneous [24].

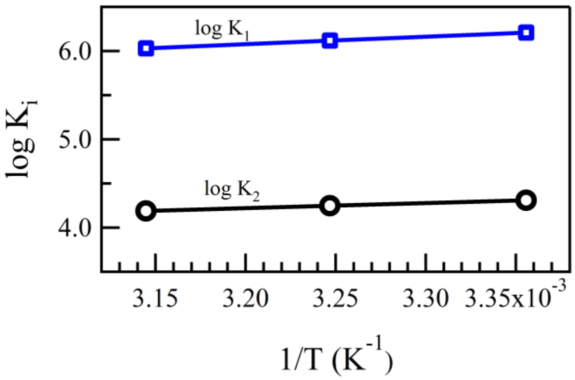

Fig. 6: Vant Hoff Plot log K_i of Re^{5+} complexes with H_2A against 1/T.

A decrease in metal-ligand stability constant (log K_i) with an increase in temperature and the negative values of enthalpy change (ΔH) for the complexation suggests that all the complexation reactions are exothermic, favorable at lower temperature and the metal-ligand binding process is enthalpy driven and metal-ligand bonds are fairly strong [25]. The negative change in entropy (ΔS) values indicated a highly solvated metal complex and indicated that the formation of these complexes was entropy favored [26, 27]. The positive values of ΔS for the complexation process of some metal ions with H_2A indicated that the complex formation is entropically favorable [28].

Table 2: Thermodynamic parameters for complexes of ligand (H_2A) in 20 % ethanol-water mixture

M^{n+}	T (K)	Free energy change (k.J.mol^{-1})		Enthalpy change (k.J.mol^{-1})		Entropy change (J.mol^{-1}.K^{-1})	
		- ΔG_1	- ΔG_2	- ΔH_1	- ΔH_2	ΔS_1	ΔS_2
Co^{2+}	298	21.97	10.68	73.78	16.31	-173.87	-18.87
	308	22.06	9.96			71.61	-20.60
	318	18.45	7.53			58.01	33.60
Pd^{2+}	298	25.39	36.27	16.33	817	30.42	94.30
	308	25.71	36.18			83.49	90.92
	318	26.00	36.13			81.76	114.07
Hg^{2+}	298	22.60	22.96	16.31	13.60	21.11	31.43
	308	22.88	22.56			74.30	29.09
	318	23.02	21.89			72.38	72.20
Re^{5+}	298	35.43	66.32	16.32	10.88	64.13	186.01
	308	36.09	66.61			117.19	180.92
	318	36.72	66.80			115.46	208.54
Th^{4+}	298	22.88	29.01	13.59	6.32	31.16	76.15
	308	23.24	29.26			75.44	74.48
	318	23.50	29.09			73.91	91.23
Ce^{3+}	298	29.73	43.76	16.32	14.53	44.97	98.09
	308	30.19	43.27			98.04	93.30
	318	30.63	42.96			96.31	137.61

4 Conclusion

The low difference between log K_i values indicates the complex formation between the metal ions and ligand occurring simultaneously. The maximum value of $\bar{n}$ was around 2, indicating the metal ion forms 1:1 and 1:2 complexes with the ligand. The metal ion solutions were very dilute (1.5×10^{-4} M), hence there was no possibility of formation of polynuclear complexes or hydrolysis of some metal ions. The large decrease in pH for the metal titration curves relative to ligand titration curve points is related to the formation of strong metal complexes. The stability constants of the formed complexes were found to be in the order $Re^{5+} > Ce^{3+} > Pd^{2+} > Th^{4+} > Hg^{2+} > Co^{2+}$. The negative change in free energy (ΔG) values indicates that both dissociation of the ligand and the complication process are spontaneous. The negative values of change in enthalpy (ΔH) for the complexation process suggest that all the complexation reactions are exothermic and favorable. The positive values of ΔS for the complexation process of some metal ions with H_2A indicated that the complex formation is entropically favorable.

References

[1] H.C. Wu, P. Thanasekaran, C.H. Tsai, Self-assembly, reorganization, and photophysical properties of silver (I)-Schiff-base molecular rectangle and polymeric array species, *Inorg. Chem.* **45**, 295–303, (2006).

[2] M. Zhao, C. Zhong, C. Stern, A.G.M. Barrett, B.M. Hoffman, Synthesis and magnetic properties comparison of M-Cu (II) and M-VO(II) Schiff base-porphyrazine complexes: what is the mechanism for spin-coupling, *J. Amer. Chem. Soc.* **127**, 9769–9775, (2005).

[3] K. Singh, M.S. Barwa, P. Tyagi, Synthesis and characterization of cobalt(II), nickel(II), copper(II) and zinc(II) complexes with Schiff base derived from 4-amino-3-mercapto-6-methyl-5-oxo-1,2,4-triazine, *European J. Medicinal Chem.* **42**, 394–402, (2007).

[4] I. Kostova, L. Saso, Advances in research of Schiff-Base metal complexes as potent antioxidants, *Current Medicinal Chem.* **20**, 4609-4632, (2013).

[5] A. Prakash, D. Adhikari, Application of Schiff bases and their metal complexes-A Review, *Int. J. Chem.Tech. Res.* **3**, 1891-1896, (2011).

[6] T. Rosu, E. Pahontu, C. Maxim, R. Georgescu, N. Stanica, A. Gulea, Some new Cu(II) complexes containing an on donor Schiff base: synthesis, characterization and antibacterial activity, *Polyhedron*, **30**, 154–162, (2011).

[7] A. Mukherjee, U. Subramanyam, V.G. Puranik, T.P. Mohandas, A. Sarkar, Pyrazole-tethered heteroditopic ligands and their transition metal complexes: synthesis, structure, and reactivity, *European J. Inorg. Chem.*, **2005**, 1254–1263, (2005).

[8] M.A.S. Goher, F.A. Mautner, M.A.M. Abu-Youssef, Synthesis, crystal structure, spectroscopic and thermal investigation of trans-diazido- tetrakis (pyrazole) manganese(II) and trans-diaquadiazidobis (3,4-dimethylpyridine) manganese (II) , Transition Metal Chem. 24, 29-34, (1999).

[9] G.G. Mohamed, M.M. Omar, A.A. Ibrahim, Biological activity studies on metal complexes of novel tridentate Schiff base ligand. Spectroscopic and thermal characterization," *European J. Medicinal Chem.* **44**, 4801–4812, 2009, (2009).

[10] G.H. Jaffery, J. Basset, J. Mendham, R.C. Denney, 1978. Vogel's Textbook of Quantitative Chemical Analysis 5th edition, Long man group, UK Limited.

[11] B. Jezowska-Trzebiatowska, S. Wajda, M. Baluka, Structure and properties of technetium and rhenium compounds of the type $[Me^{V}OX_5]^{2-}$. *J. Structural Chem.* **8**, 460–462, (1966).

[12] A.T. Mubarak, A.S. Al-Shihri, H.M. Nassef, A.A. El-Bindary, Potentiometric and thermodynamic studies of vanillin and its metal complexes, *J.Chem. & Engineering Data*, **55**, 5539–5542, (2010).

[13] R.G. Bates, M. Paabo, R.A. Robinson, Interpretation of pH measurements in alcohol-water solvents, *J. Physical Chem*. **67**, 1833–1838, (1963).

[14] H.M. Irving, M.G. Miles, L.D. Pettit, A study of some problems in determining the stoicheiometric proton dissociation constants of complexes by potentiometric titrations using a glass electrode, *Analytica Chimica Acta*. **38**, 475–488, (1967).

[15] M. Calvin, K.W. Wilson, Stability of chelate compounds. *J. Am. Chem. Soc.* **67**, 2003–2007, (1945).

[16] H. Irving, H. S. Rossotti, Methods for computing successive stability constants from experimental formation curves, *J. Chem. Soc.* (Resumed) (0): 3397-3405.

[17] A.S. Al-Shihri, A.M. Abdelhady, A.A. El-Bindary, Potentiometric and thermodynamic studies of 2-acrylamidosulfacetamide and its metal complexes in monomeric and polymeric forms, *Chem. Pap.* **58**, 155-159, (2004).

[18] A.A. El-Bindary, A.Z. El-Sonbati, E.H. El-Mosalamy, R.M. Ahmed, Potentiometric and thermodynamic studies of azosulfonamide drugs X, *Chem. Pap.* **57**, 255–258, (2003).

[19] A. Golcu, M. Tumer, H. Demirelli, R.A. Wheatley, Cd (II) and Cu (II) complexes of polydentate Schiff base ligands: Synthesis, characterisation, properties and biological activity, *Inorg. Chim. Acta.* **358**, 1785–1797, (2005).

[20] M.M. Khalil, A.M. Radalla, A.G. Mohamed, Potentiometric investigation on complexation of divalent transition metal ions with some zwitterionic buffers and triazoles, *J. Chem.& Engineering Data*, **54**, 3261–3272, (2009).

[21] L.T. Prakash, A.H. Gouri, T.N. Sharanappa, Kinetics of oxidation of antimony (III) by vanadium (V) in aqueous hydrochloric acid media. *J. Chem. Soc. Dalton Trans.* **22**, 3623–3626, (1995).

[22] V.D. Athawale, V. Lele, Stability constants and thermodynamic parameters of complexes of lanthanide ions and (±) norvaline, *J. Chem. & Engineering Data*, **41,** 1015–1019, (1996).

[23] C. Bissantz, B. Kuhn, M. Stahl, A medicinal chemist's guide to molecular interactions. *J. Med. Chem.* **53**, 5061–5084, (2010).

[24] M.F. El-Sherbiny, Potentiometric and thermodynamic studies of 2-thioxothiazolidin-4-one and its metal complexes. *Chem. Pap.* **59**, 332–335, (2005).

[25] Y. Sharmeli, R. Lonibala, Thermodynamics of the complexation of N-(pyridin-2-ylmethylene) isonicotinohydrazide with lighter lanthanides. *J. Chem. Eng. Data.* **54**, 28–34, (2009).

[26] A.A. Al-Sarawy, A.A. El-Bindary, A.Z. El-Sonbati, M.M. Mokpel, Potentiometric and thermodynamic studies of azosulfonamide drugs. *Polish J. Chem.* **80**, 289–295, (2006).

[27] D. Yang, L.E. Kay, Contributions to conformational entropy arising from bond vector fluctuations measured from NMR-derived order parameters: application to protein folding. *J. Mol. Biol.* **263**, 369-382, (1996).

[28] R.S. Saxena, K.C. Gupta, M.L. Mittal, Potentiometric and conductometric studies on the composition and stability of zinc complexes of thiomalic acid. *Canad. J. Chem.* **46**, 311–315, (1968).