



Potentiometric Studies of 3-Amino-5-Methyl Isoxazole Schiff Bases and Their Metal Complexes In Solution

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ABSTRACT

The acid-base equilibria of Schiff bases derived from 3-amino-5-methyl isoxazole and substituted salicylaldehydes, 2-hydroxy naphthaldehyde, 8-nitro-2-hydroxy naphthaldehyde, and their Co(II), Ni(II), Cu(II) and Zn(II) complex formation equilibria were investigated by potentiometrically in aqua organic media at 303K and 0.1M KNO₃. The data from the potentiometric titrations were evaluated by means of the BEST computer program. The effect of temperature and ionic strength on dissociation constant and stability constants of their complexes were studied. The corresponding thermodynamic parameter ΔG , ΔH and ΔS were calculated and discussed. The order of stability was found to be Co (II) < Ni (II) < Cu (II) > Zn (II). The concentration profile distribution of the complexes in solution has been evaluated.

Keywords: Schiff bases, potentiometry, protonation constant, formation constant, thermodynamic parameters.

1. INTRODUCTION

A Large number of heterocyclic compounds containing the isoxazole ring are associated with diverse pharmacological properties, such as anti-inflammatory, antifungal, antimicrobial activity [1-9]. Some synthetic isoxazole Schiff bases also exhibit a range of biological activities. Schiff base ligands are able to coordinate with many different metals and stabilize them in various oxidation states [10]. Azomethine and their transition metal complexes have remarkable potential for inhibiting growth of various pathogenic micro organisms [11]. It has been found that metal-ligand complexes of some isoxazole Schiff base ligands have increased biological activity [12].

Generally Schiff bases have the highest degree of hydrolysis at pH-5 and the solubility in water is also highest at this pH. Stability constants of Cu⁺², Ni⁺², Co⁺², Zn⁺², Cd⁺² and Mg⁺² with N-(3-phenyl salicylidene)-4-bromo aniline and other substituted anilines have been determined potentiometrically at 303 K and at 0.1 M ionic strength and reported earlier. The biological activity of these ligands and their complexes also studied [13].

Recently, metal complexes of Schiff base containing isoxazole ligands have been studied in the solid state by spectral characterization [14-15]. However not much of work has been carried out on the chelating tendencies of 3-(2-hydroxybenzylideneamino)-5-methyl isoxazole and its substituted derivatives with different bivalent transition metal

ions. In the present study we investigated the complexation behavior between various isoxazole Schiff bases such as MMIIMP, MIIMN, NMIIMN and AMIIMP with the bivalent metal ions such as Co (II), Ni (II), Cu (II) and Zn (II) in solution. The extent of complexation was evaluated by measuring the stability constants.

2. MATERIALS AND METHODS

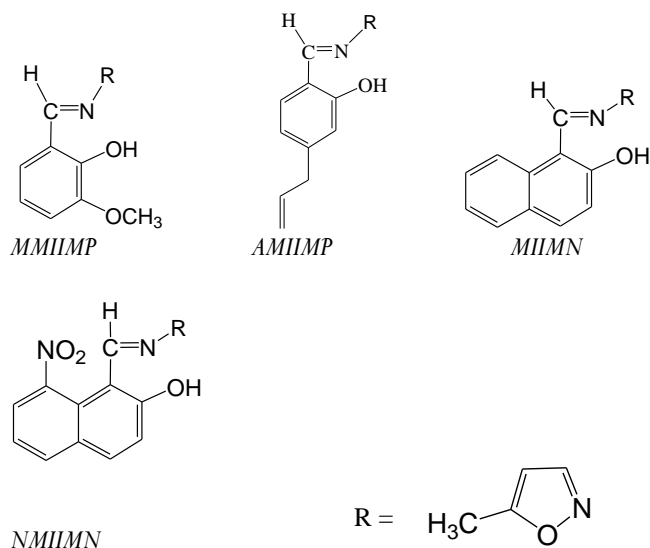
The ligand solutions were prepared by dissolving a known weight of the solid in methanol (Analar). Metal ion solutions were prepared from metal nitrates samples (BDH) in double distilled water and standardized with EDTA. Solutions of HNO₃ and 1M KNO₃ was also prepared in doubly distilled water. A carbonate free potassium hydroxide solution was used as the titrant and standardized against Analar Potassium hydrogen phthalate.

2.1. Synthesis of ligands

MMIIMP, MIIMN, NMIIMN and AMIIMP

A mixture of 3-amino-5-methyl isoxazole (0.05 mol) and 3-methoxy-2-hydroxy benzaldehyde / 2-hydroxy naphthaldehyde / 8-nitro-2-hydroxy naphthaldehyde / 4-allyl 2-hydroxy benzaldehyde (0.05 mol) was taken in methanol and refluxed for 2 hours. The yellow colored Schiff bases obtained were recrystallized from petroleum ether and methanol. Purity of the compounds was checked by TLC. Yield: 80-85%.

2.2. Structure of ligands



2.3. Potentiometric measurements

The apparatus, general conditions and methods of calculation were the same as reported earlier [16]. The following mixtures were prepared and titrated potentiometrically against standard 0.1M KOH at three different temperatures and three different ionic strengths. For each solution mixture the volume was made up to 50 ml with double distilled water before the titration. The following titrations were carried out.

1. HNO_3 (4.0×10^{-3} M)
2. HNO_3 (4.0×10^{-3} M) + ligand (1.0×10^{-3} M)
3. HNO_3 (4.0×10^{-3} M) + ligand (1.0×10^{-3} M) + metal ion (2.0×10^{-4} M)

A constant temperature was maintained using a high precision constant temperature bath (Baheti Enterprises). The pH measurements were made using a digital Digisun Electronic model DI-707 pH meter in conjunction with a combined glass and calomel electrode. The instrument could read pH in the range of 0 -14 with a reading accuracy of ± 0.1 pH units. The pH-meter readings were corrected by Van Uitert and Hass relation [17].

3. RESULTS AND DISCUSSION

3.1. Acid dissociation constants

The average number of protons, \bar{n}_A , associated with the ligand at different pH values were calculated from the titration curves of the acid in the absence and presence of ligand. The formation curves (\bar{n}_A vs. pH) for the acid systems were

constructed and found to extend between 0 and 1 in the \bar{n}_A scale. This means that, all ligands have one dissociable proton (phenolic proton). Different computational methods were applied to evaluate the stepwise dissociation constants. The average values obtained are listed in Table 1. The data in the table 1 reveals that the order of dissociation constants with respect to the ligands was found to be MMIIMP > MIIMN > NMIIMN > AMIIMP. The higher stabilities in the MMIIMP are attributed to electron donating capacity of methoxyl group at 3- position. The lower stabilities in the case of AMIIMP are due allyl substitution at 4- position in the ring. The lower pKa value of NMIIMN compared to MIIMN is due to electron withdrawing capacity of nitro group at 8-position in nitro naphthal. This order is good accordance with the basicity of the ligands.

Table 1: Dissociation constant and stepwise stability constants of metal complexes at 303K and 0.1M ionic strength in aqueous organic medium

Ligand/ Metal ion	pKa	Stability Constants		Standard Deviation
		Log K ₁	Log K ₂	
MMIIMP	8.38			
Co(II)		3.46	3.36	± 0.03
Ni(II)		4.72	3.75	± 0.02
Cu(II)		6.08	5.42	± 0.03
Zn(II)		4.26	3.92	± 0.02
MIIMN	7.23			
Co(II)		3.44	3.32	± 0.04
Ni(II)		4.32	3.69	± 0.02
Cu(II)		6.33	5.81	± 0.03
Zn(II)		3.90	3.50	± 0.02
NMIIMN	7.12			
Co(II)		3.40	3.26	± 0.03
Ni(II)		4.09	3.55	± 0.03
Cu(II)		5.51	3.91	± 0.04
Zn(II)		3.72	3.30	± 0.03
AMIIMP	6.86			
Co(II)		3.30	3.22	± 0.02
Ni(II)		3.90	3.41	± 0.03
Cu(II)		5.05	3.83	± 0.04
Zn(II)		3.55	3.24	± 0.03

3.2. Metal ligand stability constants

The formation curves for the metal complexes were obtained by plotting the average number of the ligand attached per metal ion (\bar{n}) vs. the free ligand exponent (pL), according to the Irving and Rossetti. [18]. These curves were analyzed and the successive stability constants were determined using different computational methods [19]. The representative potentiometric titration curves of MMIIMP and metals binary system in aqueous medium at 303K and 0.1M ionic strength are shown in Fig.1. The stability constants $\log K_1$ and $\log K_2$ are given in Table -1. From the data it is found that:

- The maximum value of \bar{n} was about two indicating the formation of 1:1 and 1:2 (metal: ligand) complexes only [20].
- The order of stability constants of metal complexes with respect to Schiff bases was found to be MMIIMP > MIIMN > NMIIMN > AMIIMP. This order was in accordance with the basicity order of Schiff bases.
- The order of the stability constants of metal complexes with respect to Schiff bases was found to be Co(II) < Ni(II) < Cu(II) > Zn(II), which is also in agreement with Irving and Williams order. 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 ability of the metal ion [21] and the crystal field stabilization energies [22].

Table 2: Stability constants and thermodynamic parameters of MMIIMP with transition metal ions at 0.1 M. KNO₃

Metal ion	Temp (K)	pKa	Log K ₁	Log K ₂	- Δ G kJ/mole	- Δ H kJ/mole	Δ S J/mole
Co(II)	303 K	8.45	6.86	6.01	39.80	30.87	-294.73
	313 K	8.26	6.79	5.87			
	323 K	7.12	6.69	5.84			
Ni(II)	303 K	8.45	6.93	6.08	40.20	21.79	-607.59
	313 K	8.26	6.81	5.92			
	323 K	7.12	6.77	5.90			
Cu(II)	303 K	8.45	7.41	6.42	42.98	23.61	-639.27
	313 K	8.26	7.28	6.21			
	323 K	7.12	7.22	6.17			
Zn(II)	303 K	8.45	5.61	4.66	32.54	34.50	64.69
	313 K	8.26	5.42	4.40			
	323 K	7.12	5.40	4.37			

It is evident from the data that dissociation constants and the stability constants decrease with rise in temperature, indicating that the formation equilibrium is exothermic in nature. The data in table 2 reveals that, all the metal complexes are accompanied by negative enthalpy (ΔH) changes suggesting that the metal-ligand bonds are fairly strong.

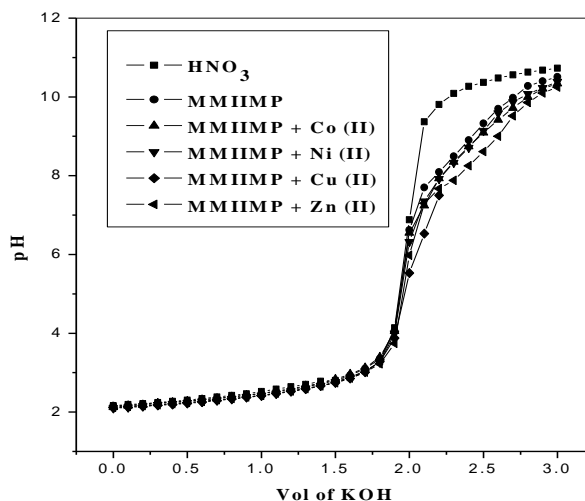
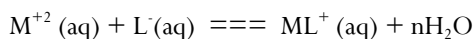


Fig. 1: pH titration curves of MMIIMP and metals binary system in aqueous medium at 303K and 0.1M ionic strength

3.3. Effect of temperature

The dissociation constant and the complexation equilibrium of MMIIMP, MIIMN, NMIIMN, and AMIIMP with bivalent metal ions have been studied at three different temperatures 303, 313 and 323K at constant 0.1M KNO₃ ionic strength and are given in Table 2. The slope of the plot ($\log K$ vs. $1/T$) was utilized to evaluate the enthalpy change (ΔH) for the complexation process. From Gibbs energy change (ΔG) and (ΔH) values one can deduce the entropy change (ΔS) using the well known relationships. The thermodynamic parameters ΔH , ΔG and ΔS were presented in Table 2.

Positive entropy changes accompanying a given reaction are due to the release of bound water molecules from the metal chelates. During formation of metal chelates, water molecules from the primary hydration sphere of the metal ion are displaced by the chelating ligand. Thus there is an increase in the number of particles in the system i.e., randomness of the system increases as shown in the following equation.



For some complexes negative entropy values are observed, which indicates that the complex largely retains the primary solvation sphere of the metal ion [23]. Williams has pointed out that usually a high entropy value is associated with the combination of positive and negative ions combination of a metal ion with a negatively charged ligand always involves displacement of water molecules, which then become part of the solvent [24]. Since the water molecules bound to the metal ions are highly distorted and oriented, the entropy is low. Large negative entropy values are associated with large negative values of enthalpy. For this, the solvent and ligand field –central ion interactions [25] and also π interactions [26] may be possible contributing factors.

3.4. Effect of ionic strength

The effect of variation of ionic strength on the stabilities of transition metal complexes of MMIIMP, MIIMN, NMIIMN and AMIIMP has been determined. For this purpose the proton ligand and metal-ligand stability constants of these ligands have been evaluated at three different ionic strengths (0.1, 0.15 and 0.2M) using potassium nitrate as a supporting electrolyte at constant temperature (303K). For some representative systems the data is summarized in Table 3 are given. From the data it is observed that the dissociation constant and stability constants decrease with raise in ionic strength.

Table 3: Stepwise stability constants of metal complexes of MMIIMP at different ionic strength at 303K temperature

Metal ion	Stability constants	Ionic strength		
		0.1M	0.15M	0.2M
H ⁺	pKa	8.45	8.36	8.32
Co(II)	Log K ₁	6.86	6.82	6.80
	Log K ₂	6.01	5.94	5.91
Ni(II)	Log K ₁	6.93	6.89	6.84
	Log K ₂	6.08	5.96	5.86
Cu(II)	Log K ₁	7.41	7.36	7.32
	Log K ₂	6.43	6.32	6.27
Zn(II)	Log K ₁	5.61	5.54	5.48
	Log K ₂	4.66	4.52	4.50

Log K values were plotted against square root of μ as per the Debye-Huckel equation. The plots of $\log K_1$ vs $\sqrt{\mu}$ for all systems were found to be linear and the slope values observed in all cases were less than the values expected on the basis of Debye-Huckel equation. Indicate that the ligands in the present

investigation interact both in dissociated and undissociated forms and thus various species co-exist in the solution [27].

3.5. Distribution of complex species in solution

In order to know the significance of metal ligand formation constants, attempts have made to see the distribution of various metal ligand species in solution with variation of pH. The degree of formation of each species involved in the complexation reaction has been calculated by using computer program BEST due to Martell and Moitekaitis in BASIC on a Microprocessor or MPF II computer. As a representative case, the distribution diagrams for some bivalent metal complexes are illustrated in Fig.2.

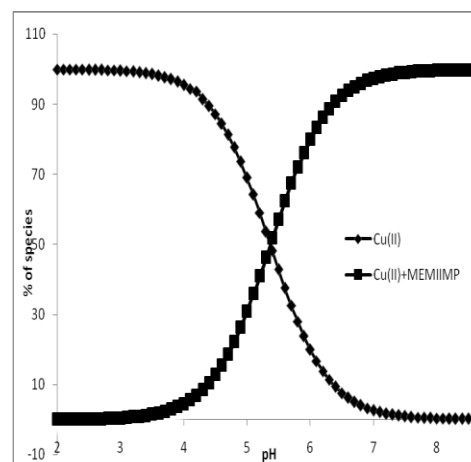


Fig.2. Species distribution curves for Cu (II)-MMIIMP binary system

The distribution curves for bivalent metal complexes show the existence of M (II), HL and L species in solution. For clarity the existence of HL and L species have been avoided in the diagram. The amount of M (II) ion decreases while the MA species increases monotonically as the pH increases. Both the M (II) ion and MA distribution curves cross-over at certain pH that corresponds to the 50% formation of the complex which varies from metal to metal and also from ligand to ligand. Similarly the pH of maximum complex formation also varies from metal to metal and ligand to ligand.

4. CONCLUSION

In conclusion, the order of stability constants of binary with respect to ligands at different temperatures and different ionic strengths was found to be MMIIMP > MIIMN > NMIIMN > AMIIMP. The order of stability constants with respect to metal ions is Co (II) < Ni (II) < Cu (II) > Zn (II) which is in accordance with Irving- Williams natural order.

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