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# THERMODYNAMIC BEHAVIOUR AND STABILITY CONSTANTS OF $UO_2^{2+}$, $Fe^{3+}$, $Cu^{2+}$, $Ni^{2+}$, $Co^{2+}$, $Fe^{2+}$, $Mn^{2+}$ AND $Cd^{2+}$ METAL COMPLEXES OF 3-METHYL-4-(2-HYDROXYPHENYL)-HYDRAZONO-2-ISOXAZOLIN-5-ONE

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Pharmaceuticals [1] are known to interact with metal ions and this has led to the study of a large number of pharmaceuticals as ligands. Further, transition metal complexes have been reported to exhibit considerable antibacterial activity [2]. A number of isoxazoles and isoxazolines have been known to have hypoglycemic activity [3,4]. Garg and Singh [5] have synthesized a few 3-methyl-4-arylhydrazono-2-isoxazolin-5-ones. Recently, we have commenced physico-chemical studies on some biologically important compounds [6–8] and here report on the potentiometric behaviour of 3-methyl-4-(2-hydroxyphenyl)hydrazono-2-isoxazolin-5-one (MHHI) and its $UO_2^{2+}$, $Cu^{2+}$, $Fe^{3+}$, $Ni^{2+}$, $Co^{2+}$, $Fe^{2+}$, $Cd^{2+}$ and $Mn^{2+}$ chelates, using the method of Irving and Rossotti [9,10].

## EXPERIMENTAL

The metal salt (nitrate or chlorides) solution was prepared in double-distilled water. All chemicals used were of AR quality. The ligand MHHI was synthesized using the method of Garg and Singh [5]. An ECIL expanded scale pH meter having an accuracy of ±0.02 pH unit was used. The pH meter readings were corrected using the method of Van Uitert and Haas [11]. The titrations were carried out in an inert atmosphere by bubbling nitrogen through the solution. The ionic strength $\mu = 0.1$ M was kept constant using sodium perchlorate. The ligand-to-metal ratio was kept at 4:1 for all the titrations. The following set of solutions were titrated against 0.1 M sodium hydroxide at 25, 30 and 35 ± 0.1°C.

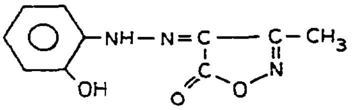
(1) 4.0 ml $HClO_4$ (0.04 M) + 4.0 ml $NaClO_4$ (1.0 M) + 30 ml methanol + 2.0 ml water;

(2) 4.0 ml $HClO_4$ (0.04 M) + 4.0 ml $NaClO_4$ (1.0 M) + 20 ml ligand (0.002 M) + 2.0 ml water + 10 ml methanol;

(3) 4.0 ml $HClO_4$ (0.04 M) + 4.0 ml $NaClO_4$ (1.0 M) + 10 ml methanol + 20 ml ligand (0.002 M) + 1.0 ml (0.01 M) metal salt solution + 1.0 ml water.

## RESULTS AND DISCUSSION

The average number of protons associated with the ligand ($\bar{n}_H$) was determined from the acid, ligand titration curves employing the equation of Irving and Rossotti. It is seen from the structure [A] of the ligand that only one proton per ligand can be released on chelation. The maximum value of $\bar{n}_H$ was found to be less than one in all cases, further suggesting that one proton per ligand molecule is liberated on coordination. The protonation constant (log $K^H$) was calculated at half integral value from plots of $\bar{n}_H$ vs. pH. The proton ligand stability constants and their thermodynamic functions at different temperatures are given in Tables 1 and 2, respectively. It is seen from Table 1 that log $K^H$ decreases with increase in temperature.



A

The metal ligand titration curves are well separated from the ligand titration curves. From the titration curves of solutions (2) and (3), the average number of ligand molecules attached per metal ion ($\bar{n}$) and free ligand exponent pL were calculated. The $\bar{n}$ values were plotted against the corresponding pL values to obtain formation curves of the complexes. In all

TABLE 1

Stability constants of MHHI and its metal complexes at various temperatures ($\mu$ = 0.1 M)

| Cation | 25°C | | | 30°C | | | 35°C | | |
|---|---|---|---|---|---|---|---|---|---|
| | log $K_1$ | log $K_2$ | log $\beta_2$ | log $K_1$ | log $K_2$ | log $\beta_2$ | log $K_1$ | log $K_2$ | log $\beta_2$ |
| $H^+$ | 8 90 | | | 8.50 | | | 8.20 | | |
| $Fe^{3+}$ | 8 25 | 5.96 | 14.21 | 8.10 | 5.68 | 13.78 | 7.90 | 5.40 | 13.30 |
| $UO_2^{2+}$ | 8 04 | 5.50 | 13.54 | 7.90 | 5.30 | 13 20 | 7.70 | 5.05 | 12.75 |
| $Cu^{2+}$ | 7.85 | 5.38 | 13.23 | 7.65 | 5.21 | 12.86 | 7.45 | 4.95 | 12.40 |
| $Ni^{2+}$ | 7.20 | 5.15 | 12.35 | 7.05 | 5.00 | 12.05 | 6.90 | 4.75 | 11.65 |
| $Co^{2+}$ | 7.10 | 5.00 | 12.10 | 6.90 | 4.80 | 11.70 | 6.65 | 4.55 | 11.20 |
| $Fe^{2+}$ | 6.75 | 4.80 | 11.55 | 6.50 | 4.65 | 11.15 | 6.30 | 4.45 | 10.75 |
| $Cd^{2+}$ | 6 20 | 4.85 | 11.05 | 6.00 | 4.70 | 10.70 | 5.75 | 4.53 | 10.28 |
| $Mn^{2+}$ | 5.80 | 4.62 | 10.42 | 5.60 | 4.45 | 10.05 | 5.40 | 4.08 | 9.48 |

TABLE 2

Values of thermodynamic functions of metal complexes at 30°C

| Cation | $-\Delta G$ (kcal mole$^{-1}$) | | | $-\Delta H$ (kcal mole$^{-1}$) | | | $-\Delta S$ (kcal mole$^{-1}$) | | |
|---|---|---|---|---|---|---|---|---|---|
| | $\Delta G_1$ | $\Delta G_2$ | $\Delta G^0$ | $\Delta H_1$ | $\Delta H_2$ | $\Delta H^0$ | $\Delta S_1$ | $\Delta S_2$ | $\Delta S^0$ |
| $H^+$ | | | 11.80 | | | 29.40 | | | 135.98 |
| $Fe^{3+}$ | 11.22 | 7.88 | 19.10 | 14.70 | 23.52 | 38.22 | 85.5 | 103.63 | 189.17 |
| $UO_2^{2+}$ | 10.95 | 7.35 | 18.30 | 14.28 | 18.90 | 33.18 | 83.26 | 86.62 | 169.70 |
| $Cu^{2+}$ | 10.60 | 7.20 | 17.80 | 16.80 | 18.06 | 34.86 | 90.42 | 83.36 | 173.80 |
| $Ni^{2+}$ | 9.78 | 6.92 | 16.70 | 12.60 | 16.80 | 29.40 | 73.86 | 78.28 | 152.15 |
| $Co^{2+}$ | 9.56 | 6.66 | 16.22 | 18.90 | 18.90 | 37.80 | 93.92 | 84.35 | 178.28 |
| $Fe^{2+}$ | 9.00 | 6.46 | 15.46 | 18.90 | 13.44 | 33.60 | 92.07 | 65.68 | 161.92 |
| $Cd^{2+}$ | 8.30 | 6.52 | 14.82 | 18.90 | 13.45 | 32.64 | 89.76 | 66.00 | 156.62 |
| $Mn^{2+}$ | 7.76 | 6.15 | 13.91 | 16.80 | 22.68 | 39.48 | 81.06 | 95.15 | 176.20 |

cases the values of $\bar{n}$ were found to be less than two, suggesting the formation of 1:1 and 1:2 complexes. The values of log $K_1$ and log $K_2$ were calculated from the formation curves using Bjeerum half $\bar{n}$ and least squares methods. The values reported in Table 1 are the means of the two methods. The metal ligand stability data log $K_1$, log $K_2$ and log $\beta_2$ at various temperatures are reported in Table 1. The order of stability for these complexes is found to be $Fe^{3+} > UO_2^{2+} > Cu^{2+} > Ni^{2+} > Co^{2+} > Fe^{2+} > Cd^{2+} > Mn^{2+}$. The stability order $Cu^{2+} > Ni^{2+} > Co^{2+} > Fe^{2+} > Cd^{2+} > Mn^{2+}$ is in accordance with the Irving–Williams [12] rule. The order $UO_2^{2+} > Cu^{2+}$ is in accordance with the greater affinity of $UO_2^{2+}$ for $O^-$–$O^-$ donors than the $Cu^{2+}$ [13]. The higher stability of the $Fe^{3+}$ complex as compared to that of divalent metal complexes is expected on the basis of the charge and ionization potential of the metal ions. The separation factor between the first and second stability constants is well within the expected range and the absence of a high value implies little steric hindrance to the addition of a second chelate group. In the case of chelating agents having bulky substituents near the donor atoms, such steric hindrance causes unusually high $K_1/K_2$. For an anionic ligand, coulombic attraction is greater for $M^{2+}$ compared to $MA^+$. As such, log $K_1$ − log $K_2$ is usually positive. Table 1 shows that for all the systems studied here log $K_1$ − log $K_2$ is positive.

The free energy, entropy and enthalpy changes have been calculated using standard relations. Values of $\Delta H$ were obtained from the slope of log $K$ vs. $1/T$ plots: wherever a linear plot was not obtained, a tangent was drawn to obtain the slope which was taken equivalent to $-\Delta H/4.57$. These thermodynamic values are given in Table 2. The error in $\Delta G^0$ and $\Delta H^0$ values lies in the range 0.4–0.8 kcal mole$^{-1}$ and that for $\Delta S^0$ is 0.8 cal deg$^{-1}$ mole$^{-1}$. It is evident from Table 2 that complex formation is exothermic and explains why the formation constants have higher values at lower temperatures.

The net entropy changes associated with the metal chelate formation can be attributed to a combination of the following:

(i) increase in $\Delta S^0$ values owing to the release of water of hydration;

(ii) decrease in entropy of translation on formation of one chelate from two species; and

(iii) decrease in configuration entropy of the ligand on complex formation.

For reactions with charged donor atoms, entropy changes associated with (i) predominate because of neutralization of charge on the metal ion; this results in a greater release of water of hydration and hence greater disorder of the partially immobilized and oriented water dipoles. The negative values of both enthalpy and entropy encountered here suggest that the enthalpy term is favourable and the entropy term is unfavourable for chelate formation. This may be due to substitution of solvent molecules attached to the metal ion by the ligand, resulting in the release of these solvent molecules.

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## REFERENCES

1 F. Lembeck and F.F. Sewing, Pharmacological Facts and Figures, Springer-Verlag, Berlin, 1970.
2 D.R. Williams, Chem. Rev , 72 (1972) 203
3 K.H. Saunders, J. Chem. Soc., 117 (1920) 1264.
4 H.G. Garg, J. Org. Chem., 27 (1962) 1045.
5 H.G. Garg and P.P. Singh, J. Med. Chem., 13 (1970) 1250.
6 A.K. Jain, R.N. Goyal, Rajeev Jain and D.D. Agarwal, J. Electroanal Chem., 93 (1978) 81.
7 R.N. Goyal, A.K. Jain and D.D. Agarwal, J. Electroanal. Chem., 93 (1978) 145.
8 A.K. Jain, R.N. Goyal and D.D. Agarwal, J. Inorg. Nucl. Chem., 43 (1981) 2005
9 H. Irving and H.S. Rossotti, J. Chem. Soc., (1954) 2904.
10 H. Irving and H.S. Rossotti, J. Chem. Soc., (1953) 3397.
11 L.G. Van Uitert and C.G. Haas, J. Am. Chem. Soc , 75 (1953) 451.
12 H. Irving and R.J.P. Williams, J. Chem Soc., (1953) 3192.
13 B.E. Bryant and W.C. Fernelius, J. Am. Chem. Soc , 76 (1954) 5351.