

POTENTIOMETRIC STUDY OF THE INTERACTION OF HYPOXANTHINE WITH SOME TRANSITION METALS

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(Received 7 September 1978)

ABSTRACT

The proton ligand stability constants of hypoxanthine and the stepwise stability constants of its complexes with manganese(II), cobalt(II), nickel(II) and copper(II) have been studied in aqueous solution at 25 and 45°C at an ionic strength of $\mu = 0.1$ M, but at 35°C with $\mu = 0.05, 0.10$ and 0.20 M. Thermodynamic stability constants have been obtained at 35°C by extrapolation to zero ionic strength. These are in the order $Mn^{2+} < Co^{2+} < Ni^{2+} < Cu^{2+}$. The free energy changes (ΔG), enthalpy changes (ΔH) and entropy changes (ΔS) associated with the formation of complexes have been calculated.

INTRODUCTION

Investigations are now taking place in order to understand the various aspects of the most significant role of metal ions in biological and metabolic processes. The metal ions, namely Mn(II), Co(II), Ni(II) and Cu(II), are all important catalysts in a variety of enzymic reactions such as group transfer, oxidation–reduction and hydrolytic processes¹. The best known biological function of Co(II) is its intimate involvement in the coenzymes related to vitamin B₁₂^{2–6}. Copper occurs in both plants and animals. A number of copper-containing proteins and enzymes have been isolated^{7, 8}. These include blue hemocyanin and ceruloplasmin. The former plays the role of an oxygen carrier in the respiratory processes while the latter is an important oxidase. The interaction of these metal ions with hypoxanthine in aqueous systems is of great interest, especially the stability constants of their complexes. Reinert and Weiss⁹ determined the stability constants of complexes of copper(II) with hypoxanthine and related compounds potentiometrically. No systematic work on these interactions is available in literature. In the present paper practical stability constants and thermodynamic parameters for the complexes of Mn(II), Co(II), Ni(II) and Cu(II) with hypoxanthine have been determined using the Bjerrum¹⁰ and Calvin and Wilson¹¹ pH titration technique as adopted by Irving and Rossotti¹².

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EXPERIMENTAL

Materials

The aqueous solutions of Mn(II), Co(II), Ni(II) and Cu(II) were obtained from BDH "Analar" grade chlorides and standardized. The strength of each solution was reduced to 0.01 M by dilution. An aqueous solution of perchloric acid (Analar, BDH) was prepared and standardized against a standard alkali to obtain 0.05 M solution of the acid. The solution of hypoxanthine (0.05 M) was prepared in water. Standard solution of KClO_4 was used to maintain the ionic strength. Standard carbonate-free potassium hydroxide solution (0.1 M) was used for titrations. The double-glass-distilled water free from carbon dioxide was used for preparing the solutions.

Apparatus

A "Systronic" pH meter model 322-1, with glass and calomel electrode assembly was used for pH measurements. Before starting the experiments the pH meter was calibrated with (BDH) buffer solutions (pH 4 and 7). SICO (Calcutta) water thermostat, type TBS, was used to maintain the temperature of the reaction mixtures at 25 ± 0.1 , 35 ± 0.1 and $45 \pm 0.1^\circ\text{C}$.

Procedure

The following solutions (initial total volume 20 ml) were titrated against the standard alkali solution (0.1 M KOH):

- (A) 2.0 ml 0.05 M HClO_4 ;
- (B) A + 2.0 ml 0.05 M ligand;
- (C) B + 2.0 ml 0.01 M metal ion.

All titrations were performed after adjusting to the proper ionic strength and maintaining at the given temperatures in a constant temperature water thermostat. The calculated quantities of 1 M potassium perchlorate solution were employed to establish the ionic strengths (0.05, 0.10, 0.20 M). The plots of pH vs. the volume of alkali used to reach it were prepared. The shapes of the titration curves were as usual.

RESULTS AND DISCUSSION

Proton-ligand stability constant

The average number of protons associated with ligand (\bar{n}_A) was calculated by the relationship

$$\bar{n}_A = Y - \frac{(V'' - V')(N^0 + E^0)}{(V^0 + V')T_L}$$

where Y is the number of dissociable protons. V' and V'' are the volumes of alkali added in the acid and the ligand titrations to reach the same pH. V^0 , N^0 and E^0 denote

the total initial volume of titration mixture, concentration of alkali and concentration of acid, respectively. T_L is total concentration of the ligand.

The formation curves were plotted between \bar{n}_A and pH. The equation

$$\log k_1^H = \text{pH} + \log \frac{\bar{n}_A}{1 - \bar{n}_A}$$

was used to obtain $\log k_1^H$ by plotting pH against $\log(1 - \bar{n}_A)/\bar{n}_A$. The $\log k_1^H$ values obtained at various ionic strengths and at different temperatures are given in Table 1. The value of $\log k_1^H$ decreases with rise of temperature and also with the increase of ionic strength.

Metal-ligand stability constants

The average number (\bar{n}) of the ligands bound per metal ion and the free ligand concentration $pL = -\log_{10} L$ were calculated from the titration curves by the following relationships of Irving and Rossotti^{1,2}.

$$\bar{n} = \frac{(V''' - V'')(N^0 + E^0)}{(V^0 + V''') \bar{n}_A T_M}$$

$$pL = \log \left[\frac{\sum_{n=0}^{n=j} \beta_n^H \left(\frac{1}{\text{antilog } B} \right)^n (V^0 + V''')}{T_L - \bar{n} T_M V^0} \right]$$

where all the terms have the usual meanings. The formation curves were obtained by plotting \bar{n} vs. pL . These curves at 35°C and $\mu \approx 0.05, 0.10$ and 0.2 M are shown in Fig. 1.

The value of \bar{n} , in all cases, is greater than 1 but it does not go beyond 2. This suggests the formation of 1:1 and 1:2 complexes in all the systems.

The values of stability constants were obtained from the formation curves by interpolation at half \bar{n} and linear extrapolation methods^{1,3}. The average values of the stability constants are given in Table 2. From a perusal of the data it is seen that the sequence of stabilities is $\text{Mn}^{2+} < \text{Co}^{2+} < \text{Ni}^{2+} < \text{Cu}^{2+}$.

This order of stability constants is in accordance with the Irving-Williams natural order¹⁴, $\text{Mn}^{2+} < \text{Fe}^{2+} < \text{Co}^{2+} < \text{Ni}^{2+} < \text{Cu}^{2+} > \text{Zn}^{2+}$ of the stability of complexes and some other orders^{15, 16}. This order is explained on the basis of the ratio of effective nuclear charge of the cation to its radius. As the value of $Z_{\text{eff}}/r_{\text{M}^{2+}}$ ratio increases, the stability of the complexes increases. On the other hand, it can be supported by the concept of atomic number. The stability of the complexes of bivalent metal ions increases as the atomic number increases. The comparison of the stability constants at 25, 35 and 45°C and at $\mu \approx 0.1$ M (Table 2) shows that the stability of the complexes decreases with the rise of temperature from 25 to 35°C, but the reverse trend is observed as the temperature changes from 35 to 45°C. There appears to be a possibility that new species which are more stable and have different chemical

TABLE 1

PROTON-LIGAND STABILITY CONSTANTS OF HYPOXANTHINE AT DIFFERENT TEMPERATURES AND IONIC STRENGTHS

Temperature ($^{\circ}\text{C}$)	Ionic strength	$\log k_1^H$
25	0.10	8.87
35	0.05	8.68
	0.10	8.65
	0.20	8.60
45	0.10	8.32

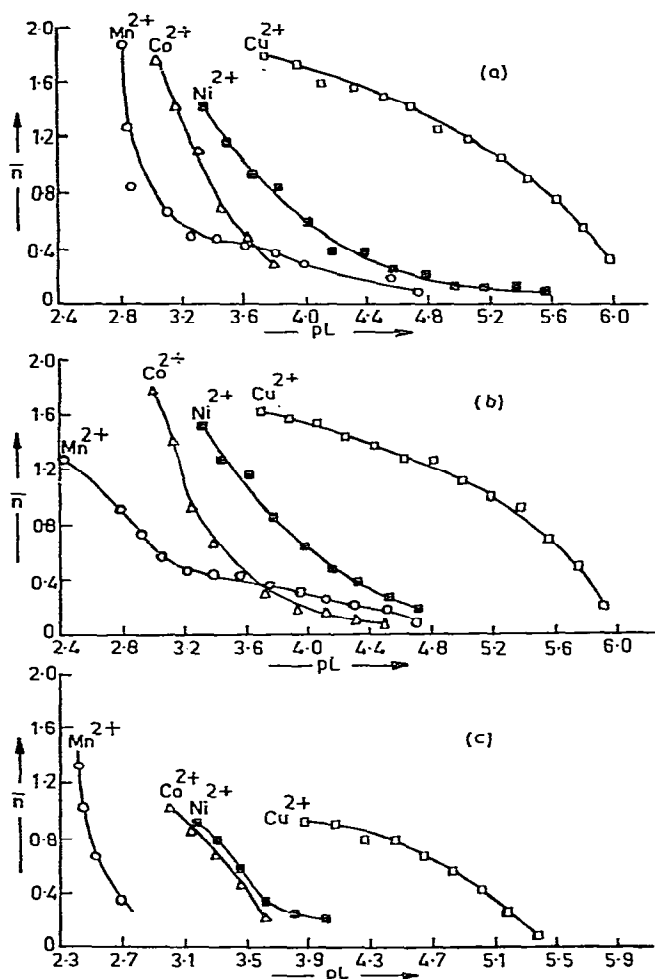


Fig. 1. Formation curves for metal-hypoxanthine complexes at 35°C and (a) $\mu = 0.05\text{ M}$, (b) $\mu = 0.10\text{ M}$, (c) $\mu = 0.20\text{ M}$.

structures are formed when temperature is increased. It is due to this that there is an increase in the values of stability constants of the complexes.

The values of $\log \beta_n$ at 35°C and at different ionic strength viz. $\mu = 0.05, 0.10$ and 0.20 M (Table 2) of the complexes decrease with the increase of ionic strength.

TABLE 2

STABILITY CONSTANTS OF METAL-HYPOXANTHINE COMPLEXES AT DIFFERENT TEMPERATURES AND IONIC STRENGTHS

Metal ion	Constant	$T = 25^{\circ}\text{C}$	$T = 35^{\circ}\text{C}$	$T = 45^{\circ}\text{C}$	$T = 35^{\circ}\text{C}$	$\mu = 0.20$
		$\mu = 0.10$	$\mu = 0.10$	$\mu = 0.10$	$\mu = 0.05$	
Mn(II)	$\log k_1$	3.50	3.00	3.42	3.34	2.56
	$\log k_2$	2.30	2.46	2.72	2.53	2.30
	$\log \beta_2$	5.80	5.46	6.14	5.87	4.86
Co(II)	$\log k_1$	3.67	3.43	3.57	3.60	3.05
	$\log k_2$	3.41	2.80	2.95	2.96	2.95
	$\log \beta_2$	7.08	6.23	6.52	6.56	5.65
Ni(II)	$\log k_1$	4.55	4.00	4.10	4.16	3.59
	$\log k_2$	3.71	2.95	3.18	3.09	2.75
	$\log \beta_2$	8.26	6.95	7.28	7.25	6.34
Cu(II)	$\log k_1$	6.22	5.59	5.77	5.76	5.54
	$\log k_2$	5.75	4.66	5.17	4.87	4.45
	$\log \beta_2$	11.97	10.25	10.94	10.63	9.99

TABLE 3

THERMODYNAMIC STABILITY CONSTANTS OF METAL-HYPOXANTHINE COMPLEXES AT 35°C

Metal ion	$\log k_1$	$\log k_2$	$\log \beta_2$
Mn(II)	4.05	2.75	6.80
Co(II)	4.25	3.25	7.50
Ni(II)	4.75	3.40	8.15
Cu(II)	6.10	5.20	11.30

This may be due to the lower activity of the metal ion for interaction with ligand with an increase in the ionic strength of the medium. Also the atmosphere about the ion is denser in salt solutions; it screens the interaction of metal ions with the ligand and thus reduces the rate of combination. This is in accordance with similar observations of Debye¹⁷ for the decrease in the stability with an increase in ionic strength of the medium.

The values of thermodynamic stability constants of the complexes were obtained by plotting $\log k_1$, $\log k_2$ and $\log \beta_2$ vs. $\sqrt{\mu}$ and extrapolating to zero ionic strength. These values are given in Table 3.

The free energy changes (ΔG) associated with the formation of the complexes were calculated by the equation $\Delta G = -2.303 RT \log K$.

These values at 25, 35 and 45°C at 0.1 M ionic strength are given in Table 4. From the data it is seen that the values of free energy change (ΔG) are negative, which suggest that the greater the negative value, the more stable is the complex. The values

TABLE 4

FREE ENERGY OF METAL-HYPOXANTHINE AT IONIC STRENGTH OF 0.1M

Metal ion	Free energy (kcal mole ⁻¹)	T (°C)		
		25	35	45
Mn(II)	-ΔG ₁	4.77	4.22	4.98
	-ΔG ₂	3.14	3.46	3.96
	-ΔG	7.91	7.68	8.94
Co(II)	-ΔG ₁	5.01	4.83	5.19
	-ΔG ₂	4.65	3.95	4.29
	-ΔG	9.66	8.78	9.48
Ni(II)	-ΔG ₁	6.21	5.63	5.97
	-ΔG ₂	5.06	4.16	4.63
	-ΔG	11.27	9.79	10.60
Cu(II)	-ΔG ₁	8.48	7.88	8.40
	-ΔG ₂	7.84	6.57	7.52
	-ΔG	16.32	14.45	15.92

TABLE 5

ENTHALPY CHANGES OF METAL-HYPOXANTHINE AT IONIC STRENGTHS OF 0.1 M

Metal ion	ΔH ₁ (kcal mole ⁻¹)	ΔH ₂ (kcal mole ⁻¹)	ΔH ^a (kcal mole ⁻¹)
Mn(II)	-21.00	6.78	-14.22
Co(II)	-10.08	-25.62	-35.70
Ni(II)	-23.10	-31.92	-55.02
Cu(II)	-26.45	-45.77	-77.22

^a ΔH = ΔH₁ + ΔH₂ calculated from formation constants at 25 and 35°C.

TABLE 6

ENTROPY OF METAL-HYPOXANTHINE COMPLEXES AT IONIC STRENGTH OF 0.1 M

Metal ion	T (°C)	ΔS ₁ (cal deg ⁻¹ mole ⁻¹)	ΔS ₂ (cal deg ⁻¹ mole ⁻¹)	ΔS (cal deg ⁻¹ mole ⁻¹)
Mn(II)	25	-54.46	33.28	-21.18
	35	-54.48	33.24	-21.24
Co(II)	25	-17.01	-70.37	-87.38
	35	-17.04	-70.35	-87.39
Ni(II)	25	-56.67	-90.13	-146.80
	35	-56.72	-90.12	-146.84
Cu(II)	25	-60.30	-127.20	-187.50
	35	-60.29	-127.27	-187.56

of ΔG in the case of Mn(II), Co(II), Ni(II) and Cu(II) complexes are in the order Mn(II) < Co(II) < Ni(II) < Cu(II). This also justifies the observed stability order as given elsewhere.

The following relationships were employed in order to obtain the enthalpy changes (ΔH) and entropy changes (ΔS) from the $\log k$ values at 25 and 35°C at $\mu = 0.1$ M ionic strength.

$$\Delta H = 2.303 R \frac{T_1 T_2}{T_2 - T_1} (\log k_{T_2} - \log k_{T_1})$$

$$\Delta G = \Delta H - T\Delta S$$

The values obtained for these functions are given in Tables 5 and 6. The enthalpy and entropy changes of complex formation are often described by using a simple electrostatic model in connection with the Frank and Evans iceberg concept. Species in water solutions are surrounded by a cloud of water molecules with a geometry different to that of bulk water. The decrease in entropy observed in solution is also associated with orientation of the water molecules around the ions with an attendant loss in freedom of motion. The complex has greater radius and reduced charge as compared to the reactants and, therefore, water molecules associated with the ions are released during the reactions. It was pointed out by Williams¹⁸ that where a high entropy of reaction is observed, the reaction usually involves a combination of positive and negative ions.

It is observed that values of ΔH and ΔS are negative for all the complexes, which indicates that only ΔH is favourable for complexation. The negative values of ΔS for Mn(II), Co(II), Ni(II) and Cu(II) suggest that the primary hydration sphere is retained and the complexes are of outer sphere, ion pair type.

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