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Kinetics of Formation of the Mixed Ligand Complex between Ni(Nitrilotriacetate)$^-$ and Imidazole [1]

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Introduction

Mixed ligand complexes attract much attention [2, 3] due to their roles in natural systems. To mimic and to improve our understanding of the selectivity observed in biological systems we have studied [1] the stability of ternary M(Nta)(A) complexes, where $M^{2+} = Co^{2+}, Ni^{2+}, Cu^{2+}$ or Zn^{2+}, Nta^{3-} = nitrilotriacetate and A = pyridine, imidazole, NH_3, CH_3COO^- or HPO_4^{2-}. These monodentate ligands are the simplest models for the corresponding ligating groups often occurring in nature; 'pyridine'-nitrogens are found *e.g.*, in pyrimidine and purine derivatives. The most fascinating aspect of this study [1] is that pyridine and imidazole form ternary complexes which are more stable than the corresponding binary ones, *i.e.* for imidazole (Im) and the mentioned M^{2+} $\Delta \log K_M = \log K^{M(Nta)}_{M(Nta)(Im)} - \log K^{M}_{M(Im)} = +0.02$ to $+0.25$; these values are significantly larger than expected on a statistical basis: $\Delta \log K_{st/oh} = -0.5$ [1]. This increased stability has been attributed to the π-accepting properties of imidazole [4] in accordance with related observations [3, 5]. In general, such enhanced stabilities are observed in combinations of a heteroaromatic N base and an O-donor such as a carboxylate, phenolate or phosphate.

To learn whether this enhanced thermodynamic stability is due to an increased rate of formation, or a reduced rate of dissociation or a combination of both we have now studied the kinetic properties of the Ni(Nta)$^-$/imidazole system. Certainly, Ni^{2+} seems of minor biological importance [6, 7], compared *e.g.* with Zn^{2+}, but it is especially suitable for kinetic studies [8], because data are available for comparisons [8–14]. Kinetic studies with the aim to evaluate the reasons for the mentioned enhanced stability of mixed ligand complexes are still very scarce [15–18] and no general conclusions can so far be drawn.

Experimental

The chemicals were the same as used previously [1]. The kinetics of complex formation between Ni(Nta)$^-$ and imidazole were studied at 563 nm using a Durrum D 150 T-jump spectrophotometer, interfaced with a Hewlett-Packard HP 9820 computer through a Datalab 901 transient recorder [19].

All solutions contained Ni^{2+} and Nta^{3-} in a 1:1 ratio (5×10^{-3} to 2.6×10^{-2} *M*); the imidazole concentration was varied from 2.6×10^{-3} to 2.6×10^{-2} *M* and the pH from 5.0 to 5.8 (I = 0.1 *M*, KNO_3). The release of protons during the complexation reaction was followed by 5×10^{-5} *M* methyl red, present in all solutions. The desired pH was adjusted to ±0.01 pH-units with small amounts of HNO_3 or KOH, and just before the T-jump measurement the solutions were filtered through Millipore filters and degassed. The temperature of the experiments was 25 ± 0.05 °C.

All solutions containing Ni^{2+}, Nta^{3-} and imidazole showed only a single relaxation effect. Blank tests with solutions containing only Ni^{2+}/Nta^{3-} or imidazole gave no relaxation signal in the time range of the instrument. The experimental points were fitted to a single exponential function, from which the relaxation time τ was obtained. The τ values used for the final calculation are the mean of at least 5 measurements with a maximum relative error of ±5%.

Large ratios of $[Ni(Nta)^-]/[Im]$ and $[H(Im)^+]/[Im]$ were used throughout to assure that only the mono-imidazole complex formed.

Results and Discussion

Complex formation between imidazole and Ni(Nta)$^-$ was studied in the pH range 5 to 5.8, where imidazole is mainly present as $H(Im)^+$. There are reports [13, 14] that $H(Im)^+$ may directly react with Ni^{2+}_{aq} at a rate 8 to 25 times smaller than that of Im itself. As Ni(Nta)$^-$ carries a negative charge the reaction with $H(Im)^+$ might become even more important, therefore reactions (1) and (2) were considered.

$$Ni(Nta)^- + Im \underset{k_{-1}}{\overset{k_{+1}}{\rightleftharpoons}} Ni(Nta)(Im)^- \tag{1}$$

$$Ni(Nta)^- + H(Im)^+ \underset{k_{-2}}{\overset{k_{+2}}{\rightleftharpoons}} Ni(Nta)(Im)^- + H^+ \tag{2}$$

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TABLE I. Rate Constants for the Complexation of Ni^{2+} and $Ni(Nta)^-$ with Imidazole and Ammonia in Aqueous Solution at 25 °C.

	Im k_{+1} (M^{-1} s^{-1})	NH_3 k_{+1} (M^{-1} s^{-1})	Im k_{-1} (s^{-1})	NH_3 k_{-1} (s^{-1})
Ni^{2+}	3.2×10^3 [13][a]	2.8×10^3 [11][d]		
	5×10^3 [9][b]	4.3×10^3 [10][a]		
	6.4×10^3 [14][a]	4.5×10^3 [12][e]		
	av.:4.9×10^3	3.9×10^3	4.9[f]	7.1[f]
$Ni(Nta)^-$	5.7×10^4 [*cf.*[c]]	4.6×10^3 [11][d]	54.4[f]	13.3[f]

[a] I = 0.1. [b] I = 0.15. [c] This work; the error limits are $\pm 0.1 \times 10^4$; I = 0.1. [d] I = 0.25. [e] I = 1.0. [f] Calculated from the listed values of k_{+1} and the stability constants given in eqns. (7) and (8).

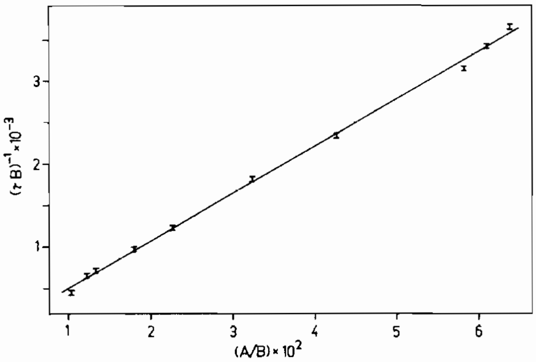

Fig. 1. Plot of $(\tau \cdot B)^{-1}$ *versus* A/B for the $Ni(Nta)^-$/imidazole system. According to eqn. (6) the 'best' straight line was calculated using the least-squares procedure: $(\tau B)^{-1} = (57.000 \pm 1.000) \cdot (A/B) + (-73 \pm 33)$. The standard deviation of the individual $(\tau B)^{-1}$ values is given by the length of the bars.

These reactions are coupled to the more rapid processes (3) and (4), where H(Ind) is the acidic form of the indicator methyl red.

$$H(Im)^+ \rightleftharpoons H^+ + Im \qquad (3)$$

$$H(Ind) \rightleftharpoons H^+ + Ind^- \qquad (4)$$

The constants for eqns. (1), (3), and (4) are $K^{Ni(Nta)}_{Ni(Nta)(Im)} = 1.05 \times 10^3\ M^{-1}$, $K^{H}_{H(Im)} = 9.12 \times 10^{-8}\ M$ [1], and $K^{H}_{H(Ind)} = 1 \times 10^{-5}\ M$ [20], respectively.

By standard techniques [15, 21] one obtains for the relaxation times eqn. (5),

$$1/\tau = k_{+1} \cdot A + k_{+2} \cdot B \qquad (5)$$

which corresponds in its rearranged form (6) to the equation of a straight line.

$$(\tau \cdot B)^{-1} = k_{+1} \cdot (A/B) + k_{+2} \qquad (6)$$

In eqns. (5) and (6) are

$$A = [Ni(Nta)^-]/(1 + \alpha) + [Im] + 1/K^{Ni(Nta)}_{Ni(Nta)(Im)}$$

$$B = \alpha[Ni(Nta)^-]/(1 + \alpha) + [H(Im)^+]$$

$$+ \{[H^+] + \alpha\beta[Ni(Nta)^-]/(1 + \alpha)\}/\{K^{Ni(Nta)}_{Ni(Nta)(Im)} K^{H}_{H(Im)}\}$$

$$\alpha = [H^+]/(K^{H}_{H(Im)} + \beta[Im])$$

$$\beta = (K^{H}_{H(Ind)} + [H^+])/(K^{H}_{H(Ind)} + [H^+] + [Ind^-])$$

The values of A and B were calculated for each pH and for the total reactant concentrations used. A plot of $(\tau \cdot B)^{-1}$ *versus* A/B results, as expected according to eqn. (6), in a straight line which has k_{+1} as slope and k_{+2} as intercept (*cf.* Fig. 1). The slope equals $(5.7 \pm 0.1) \times 10^4\ M^{-1}\ s^{-1}$ and the intercept $-73 \pm 33\ M^{-1}\ s^{-1}$, which means that $k_{+2} \cong 0$; hence, under the experimental conditions employed in this study the protonated form $H(Im)^+$ is not reactive towards $Ni(Nta)^-$.

The rate constants of equilibrium (1) are given in Table I together with related data [9–14]. The thermodynamic results [1] should also be recalled:

$$\Delta \log K_{Ni/Im} = \log K^{Ni(Nta)}_{Ni(Nta)(Im)} - \log K^{Ni}_{Ni(Im)} =$$

$$= 3.02 - 3.00 = +0.02 \qquad (7)$$

$$\Delta \log K_{Ni/NH_3} = \log K^{Ni(Nta)}_{Ni(Nta)(NH_3)} - \log K^{Ni}_{Ni(NH_3)} =$$

$$= 2.54 - 2.74 = -0.20 \qquad (8)$$

The rate constants for ligand substitution in Ni^{2+}_{aq} with NH_3 and imidazole are determined by the solvent exchange rate [8]. Indeed, the corresponding rate constants are similar (Table I) and the slightly larger stability of $Ni(Im)^{2+}$, compared with $Ni(NH_3)^{2+}$, is the result of a slightly enhanced formation and a somewhat reduced dissociation rate. However, both constants k_{+1} and k_{-1} for $Ni(Nta)(Im)^-$

are more than 10-fold enhanced*, if compared to the values of $Ni(Im)^{2+}$; hence $\Delta\log K_{Ni/Im} \cong 0$ (eqn. 7). The corresponding comparison with the NH_3 systems reveals that here the differences are much smaller. Therefore the rates of the two ternary systems are discussed more in detail:

The substitution rate of $Ni(Nta)^-_{aq}$ for imidazole is about 12 times faster than for NH_3 (Table I), while the dissociation rate is increased only by a factor of about 4. The formation rate constant k_{+1} is usually attributed to the product $K_{os} \cdot k_{solv}$, where K_{os} is the stability constant and k_{solv} the solvent exchange rate of the outer-sphere complex [8]. As one may assume that k_{solv} is similar for both substitutions, *i.e.* with NH_3 and Im, the higher value of k_{+1} for Im appears to reflect a higher stability of the outer-sphere complex. How this could result is not clear: it might be that Im is favored for a certain hydrogen bond, or that a hydrophobic interaction between ethylene groups of the coordinated Nta^{3-} and the aromatic system of Im occurs. Such hydrophobic interactions in ternary complexes are known [22] and the related stacking interactions [3] have been used to explain enhanced formation rates [8, 23]. This latter explanation is appealing: (i) it disfavors NH_3 and (ii) it would allow to rationalize also the enhanced stability of $M(Nta)(pyridine)^-$. On the other hand one must admit that this explanation cannot universally hold for all ternary systems showing an enhanced stability and containing a heteroaromatic N base and an O-donor [3–5], because, *e.g.*, ternary $M(2,2'\text{-bipyridyl})(HPO_4)$ complexes have also positive values for $\Delta \log K_M$ [24], and for these no hydrophobic interactions can be assumed.

To conclude, the relatively large stability of $Ni(Nta)(Im)^-$, compared with $Ni(Nta)(NH_3)^-$, is mainly due to an enhanced formation rate, but it remains to be seen if this holds also for the corresponding complexes of the other metal ions.

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*Taking into account the statistical factor of 2/6, k_{+1} is 30-fold enhanced.