Kinetics of complexation of nickel(II) by DL-isocitric acid

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Summary

The kinetics of complexation of Ni^{II} by isocitric acid have been studied with the stopped-flow method, at 15, 25 and 35° C, ionic strength 0.20 M ($NaClO_4$) and pH range 4.50–6.35. Under our experimental conditions, two processes are observed: the faster reaction takes place within the millisecond time range and the slower one within a few seconds.

A mechanism is proposed to account for the observed behaviour. The rapid process is associated with the complexation itself and the slow one is associated with the deprotonation. For the latter process, the activation energy was determined.

Introduction

Isocitrate complexes with the metallic ions, particularly of the transition series, play a relevant role in many enzymatic processes[1][2]. Nickel(II) complexes with hydroxypolycarboxylic acids are relevant in plant physiology[3].

Kinetic, spectroscopic and thermodynamic information regarding the interaction between divalent cations and isocitric acid is scarce. To the best of our knowledge, the only kinetic study reported in the literature involves magnasium(II). Proton nmr[5][6] and potentiometric studies[7] for cobalt(II), manganese(II), etc. have also been reported.

Isocitric acid has several coordination centres, three carboxylate groups and one OH group. The kinetics of coordination with a metal centre could be affected by steric hindrance. In such a case. the system would depart from a normal or Eigen mechanism[8-12].

The principal aim of this work was to elucidate the kinetics and mechanism of complexation between Ni^{II} and DL-isocitric acid and the nature of the interaction between the reactants.

Experimental

The kinetic study was carried on a Durrum D-110 stopped-flow apparatus at 15, 25 and 35° C.

Reagent grade $Ni(ClO_4)_2 \cdot 6H_2O$ (Fluka AG), DL-isocitric acid (Serva, Heidelberg), $NaClO_4$ (Merck) and pH indicators (all Mallinkrodt) were employed without further purification.

In all experiments the nickel ion concentration was always in an excess (ten-fold) over the carboxylic acid to maintain pseudo-first-order conditions. The concentration ranged from 0.5×10^{-2} to 2.0×10^{-2} M. The study was performed in a 4.50–6.35 pH range. The pH of each solution was adjusted to the desire value before mixing. The reported pH is that of the solution after mixing. The absorbance changes were recorded using an appropriate pH indicator added to the system (1 $\times 10^{-5}$ M); methyl red (pK = 5.0) and chlorophenol red (pK = 6) were employed. The absorbance changes were

followed at 517 nm (for methyl red) or 580 nm (for chlorophenol red). The ionic strength was 0.2 M ($NaClO_4$).

Stopped-flow measurements with reactants show two processes well resolved in time. The faster process was observed in the millisecond time range, whereas the slower one was recorded within the second time range. The processes show opposite absorbance changes.

No absorbance changes were recorded in blank experiments with nickel(II)-indicator or carboxylic acid-indicator.

To analyse the slow reaction, another set of experiments with nickel(II) and tricarballylic acid (Aldrich) was carried out. No absorbance changes were recorded under the experimental conditions employed for the nickel(II) isocitrate system.
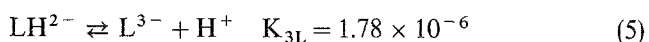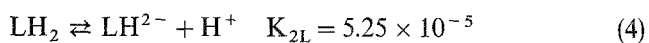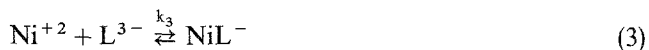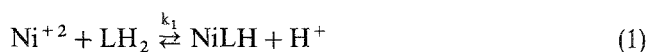
Results

The experimental conditions and the measured rate constants are listed in Table 1. These results were obtained form averaging at least five oscillographic records. The deviation was usually less than 15%.

For simplicity the fast and slow processes will be discussed separately.

Fast process

The nickel(II) and proton concentration-dependence on the observed rates associated with the fast process can be interpreted in terms of the following mechanism.

$$Ni^{+2} + LH_2 \overset{k_1}{\rightleftarrows} NiLH + H^+ \tag{1}$$

$$Ni^{+2} + LH^{2-} \overset{k_2}{\rightleftarrows} NiLH \tag{2}$$

$$Ni^{+2} + L^{3-} \overset{k_3}{\rightleftarrows} NiL^- \tag{3}$$

$$LH_2 \rightleftarrows LH^{2-} + H^+ \quad K_{2L} = 5.25 \times 10^{-5} \tag{4}$$

$$LH^{2-} \rightleftarrows L^{3-} + H^+ \quad K_{3L} = 1.78 \times 10^{-6} \tag{5}$$

LH_2^-, LH^{2-}, LH^{3-} are the different possible forms of isocitric acid; NiLH and NiL^- are the complexes. However, in the pH range where this reaction was studied the mechanism can be simplified, omitting step (1) which is inconsistent with the results.

Taking into account the complexation steps (2) and (3), the observed apparent rate constant k_{APP} can be written as (6).

$$k_{APP} = k_r + [Ni^{2+}](k_2 + k_3 \cdot K_{3L}/[H^+])/\Delta \tag{6}$$

$$k_{APP} = \ln 2/t_{1/2}$$
$$\Delta = 1 + K_{3L}/[H^+] \tag{7}$$

In Equation (6), k_r involves the reverse reactions (2) and (3); finally, the nickel concentration corresponds to the analytical one.

Regrouping terms in Equation (6), an expression of the

Table 1. Apparent rate constants for the nickel(II) and isocitric acid system.

pH	Rapid process [Ni] × 10³/mol dm³					Slow process t/°C		
	2.5	5.0	10	15	20	15	25	35
4.50		157.3						
4.55		123.6						
4.60		173.0					0.75	
4.65		203.5					0.52	
4.70							0.53	
4.75	117.3	384.4				0.25		
4.80	102.7				576.7		0.52	
4.85			314.5	432.5	432.5		0.34	
4.90				576.7	494.3	0.27	0.36	1.00
4.95	118.2		247.1			0.21	0.52	0.96
5.00		256.3	329.5	432.5	601.7		0.67	0.82
5.05				346.0		0.24	0.46	0.88
5.10	86.5				494.3		0.35	0.91
5.15		223.2			532.3	0.14	0.24	0.82
5.20			266.2	494.3	432.5	0.16	0.23	0.58
5.25	108.0	197.7				0.17	0.27	0.56
5.30				432.5	461.3	0.21	0.23	
5.35	138.2		314.5		576.7		0.26	
5.40	123.6		384.4			0.14	0.21	0.54
5.45	127.7		384.4				0.25	
5.50		247.1	329.5			0.11	0.21	
5.55							0.18	0.55
5.60	128.3			461.3	629.0	0.13	0.18	0.42
5.65							0.18	0.52
5.70			407.1	407.1		0.13	0.19	0.44
5.75	160.5	314.5			629.1	0.12	0.21	0.48
5.80			384.4		692.0	0.12	0.20	0.49
5.85		300.9		461.3		0.093	0.21	
5.90		266.2	384.4			0.11	0.18	0.61
5.95	138.4	266.2			692.0			0.47
6.00	164.7	314.5		461.3			0.094	0.47
6.05	157.3			407.1				
6.10				576.7	576.7			
6.15		329.5		432.5				
6.20				407.1	576.7		0.12	0.15
6.25			494.3					
6.30			432.5					
6.35			461.2					

form

$$F = k_2^* + k_3^*/[H^+]$$

is obtained, in which expressions (8)–(10) hold.

$$F = k_{APP}\cdot(1 + K_{3L}/[H^+]) \qquad (8)$$

$$k_2^* = k_2 + k_r/[Ni^{2+}] \qquad (9)$$

$$k_3^* = \{k_3 + k_r/[Ni^{2+}]\}K_{3L} \qquad (10)$$

A plot of F vs $1/[H^+]$ is shown in Figure 1.

The intercept (k_2^*) and the slope (k_3^*) for different total ion concentration are listed in Table 2.

The individual rate constants k_2, k_3 and k_r were obtained from plots of the intercepts and the slope versus $1/[Ni]$ are shown in Figure 2.

The values obtained were $k_2 = 2.6 \times 10^4\,mol\,dm^3\,s^{-1}$ and $k_3 = 3.4 \times 10^4\,mol\,dm^3\,s^{-1}$. The values of k_r lie between 50 and $130\,s^{-1}$.

Slow process

Under the experimental conditions reported for the fast process, a slower reaction can be detected. This process is independent of the ion and isocitric acid concentrations. However, the apparent rate constant is a function of pH only, with a linear dependence as shown in Figure 3, where apparent rate constants (k_{2APP}) measured at 15, 25 and 35° C are represented.

A possible interpretation involves deprotonation of the

Table 2. Kinetic data for observed rapid process at 25° C.

[Ni] × 10³/mol dm⁻³	k_2^* /s⁻¹	k_3^* /s⁻¹	r*
2.5	4.24 × 10⁴	0.140	0.98
5.0	3.46 × 10⁴	0.143	0.98
10.0	2.92 × 10⁴	0.093	0.99
15.0	3.17 × 10⁴	0.056	0.97
20.0	2.55 × 10⁴	0.059	0.96

*Correlation coefficient.
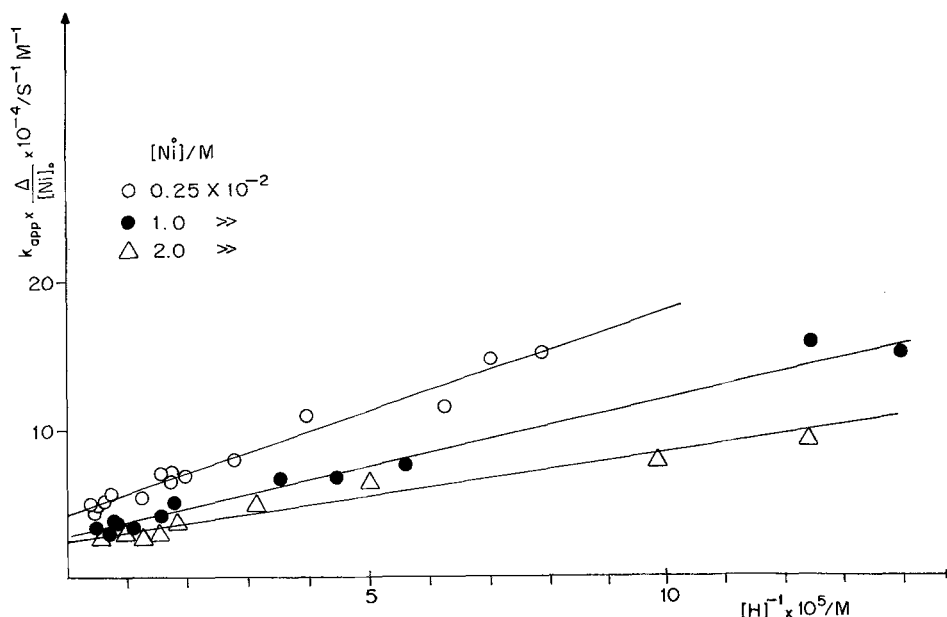
Figure 1. Rapid processes: F versus 1/[H] a 25° C.
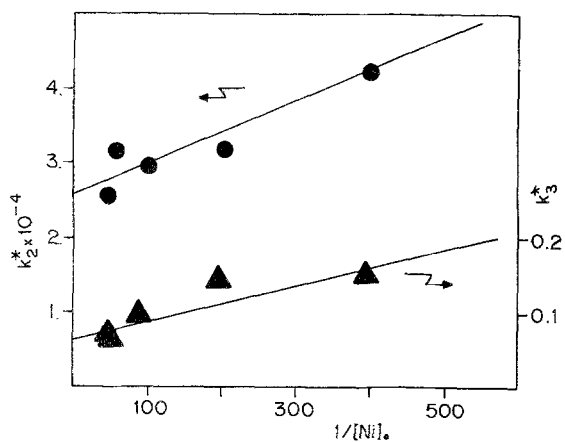
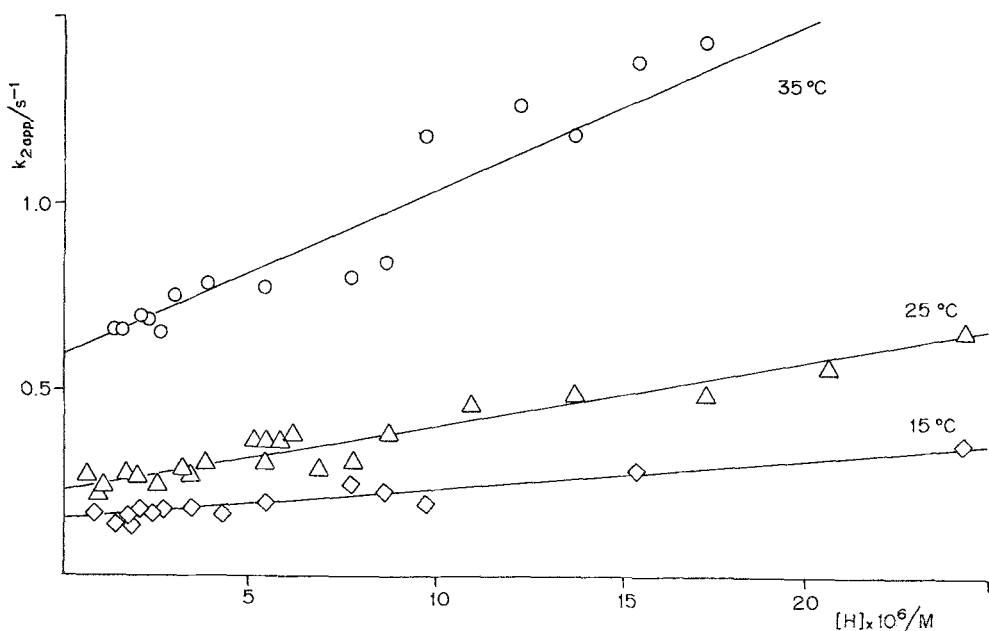**Figure 2.** Rapid processes: k_2^* and k_3^* versus $1/[Ni]_0$ at 25° C.

system under study, the rate of complex formation would be governed by the rate k_{exch} at which the water molecules leave the inner-ion hydration sphere[9-12]. This rate for Ni^II has been measured by n.m.r.[13] and its value is $3 \times 10^4 \, s^{-1}$[14]. Apart from electrostatic effects, the global rate constant should be independent of the ligand. Reactions (2) and (3) can be considered as composed of the following step-by-step mechanism.

$$[Ni(H_2O)_6]^{2+} + L_i^{n-} \rightleftharpoons [Ni(H_2O)_6|L_i]^{2-n} \quad K_0 \qquad (12)$$

$$[Ni(H_2O)_6|L_i]^{2-n} \overset{k_{exch}}{\rightleftharpoons} [Ni(H_2O)_5 - L_i]^{2-n} + H_2O \qquad (13)$$

L_i represent LH^{2-} or $L^{3-} \cdot K_0$ is the ion-pair equilibrium constant, which can be estimated by the well-known Fuoss-Bjerrum Equation[14,15]. The expected rate

Figure 3. Slow process: k_{APP} versus [H] at 15°, 25° and 35° C.

complex by interaction of the carboxylic acid OH group with the central ion.

Representing the complex forms as Ni — L and Ni = L, the following global reaction can be considered,

$$Ni - L \overset{k_4}{\underset{k_{-4}}{\rightleftharpoons}} Ni = L + H \qquad (11)$$

Here charges have been omitted for the sake of simplicity.

This simple mechanism leads to the following expression for the observed rate constant k_{2APP}.

$$k_{2APP} = k_4 + k_{-4} \cdot [H^+] \qquad (12)$$

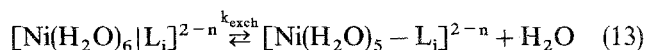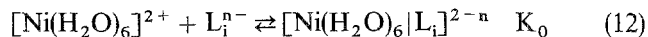The Arrhenius plots for the slope (k_{-4}) and the intercept (k_4) are shown in Figure 4.

The obtained values for the slope and the intercepts at these temperatures, together with the activation energy, are listed in Table 3.

Discussion

Fast reactions (2) and (3) describe complexation steps, where the carboxylic groups of the ligand are involved.

If the general Eigen mechanism were valid for the

constants of the Eigen mechanism are given by $K_0 \cdot k_{exch}^{(9-12)}$.

With the preceeding considerations, the expected rate constants are $2.6 \times 10^5 \, mol \, dm^{-3} \, s^{-1}$ for reaction (2) and $1.4 \times 10^6 \, mol \, dm^{-3} \, s^{-1}$ for reaction (3). The K_0 values are 8.7 and 45.7 for reactions (2) and (3), respectively.

The experimental rates differ from the expected values because of steric hindrance in the complex formation. In these complexes rings of at least six members are involved.
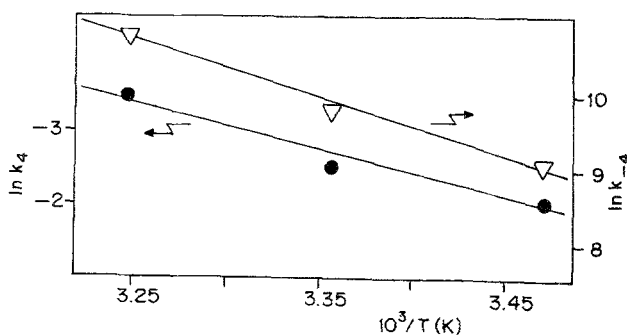

Figure 4. The temperature dependence of k_4 and k_{-4}.

Table 3. Kinetic data for slow process.

T/K	k_4 (s^{-1})	k_{-4} (mol·dm^{-3} s^{-1})	E_4 (kJ/mol)	E_{-4}
288	0.15	8.90×10^3		
298	0.23	1.75×10^4	50.0	60.7
308	0.58	4.71×10^4		

It is well known that in rings of six or even seven members steric effects are relevant[10]. For reaction (3), in addition to steric hindrance, another possible factor could be the intramolecular hydrogen bond between the OH and one carboxylic group in the isocitric acid[10–19]. Similar behaviour was reported for other ligands[20, 21]. In such a case, the global mechanism would be also controlled by intramolecular hydrogen bonds.

The slow process can be ascribed to the deprotonation of the complexes [represented by Ni–L in reaction (11)]. In this step, an OH group should be involved, leading to the formation of a bond with the metal ion. This complex form is represented by Ni = L in reaction (11).

Experiments with tricarballylic acid support these considerations. This tricarboxylic acid does not possess an OH group. In the presence of tricarballylic acid, no absorbance change is observed. This allows us to reject other possible interpretations, for if the slow reaction involved the formation of binuclear complexes, as observed for citric acid in alkaline media[22], a similar reaction with tricarballylic acid and other hydroxy-policarboxylic acids would be expected. However, no evidence of binuclear complex formation has been reported for similar reactions in this pH range[21, 23–28].

Early studies performed with nickel(II) and malate and tartrate[26, 27] show that both form six-membered chelate rings, with coordination through the carboxylic and hydroxy groups. Moreover, since the interaction of the hydroxy oxygen with the metal ion is weaker than that of the carboxylic group, a longer bond in the latter bond has been postulated. A pressure jump study of the nickel(II)-lactate complex formation[28] showed that the rate-determining step of the process is associated with the chelation of the OH group.

All the preceding considerations support the hypothesis for the slow process in the nickel ion and the isocitrate system. The activation parameters for the forward and reverse reactions, 50 and 60.7 kJ/mol, are very close to the activation energies for other ligands reported in the literature[24, 26]. These similarities imply that the rate-limiting step of the observed behaviour for isocitrate and other related ligands is governed by bond formation between the metal ion and the oxygen atom derived from the OH group.

Acknowledgments

References

[1] E. Ochiai, *Bioinorganics Chemistry. An Introduction*, Allyn Bacon Inc., Boston.

[2] R. F. Colman, *J. Biol. Chem.*, **247**, 215 (1973).

[3] U. Kersten, R. Brooke, R. Reeves and T. Kersten, *Phytochem*, **19**, 1963 (1980).

[4] M. R. Feliz and G. Ellenrieder, *Anal. Asoc. Qca. Argent.*, **61**, 265 (1972).

[5] E. J. Bastian and R. Martin, *J. Phys. Chem.*, **76**, 3073 (1972).

[6] G. Occupati and J. Pratt, *J. Inorg. Nucl. Chem.*, **36**, 1713 (1974).

[7] A. E. Mantell and R. M. Smith, *Critical Stability Constants*, Vol 3 (1977), Vol 5 First Suppet (1982), Plenum Press, New York and London.

[8] M. Eigen and K. Tamm, *Z. Elecktroch.*, **64**, 115 (1960); *ibid. idem*, **66**, 93 (1962).

[9] R. F. Gould, Ed. in M. Eigen and R. G. Wilkins, *Mechanisms of Inorganic Reactions*, In *Advances in Chemistry Series*, Vol. 49, American Chemical Society (1965).

[10] O. Edwards, Ed. in M. Kustin and J. Swinehart, *Inorganic Reaction Mechanism*, in *Progress in Inorganic Chemistry*, vol. 13, J. Wiley, New York.

[11] U. Nickel, H. Hoffmann and W. Jaenicke, *Ber Bunsenges. Physik. Chem.*, **72**, 526 (1968).

[12] S. Petrucci, *Ionic Interactions*, vol. 2, Academic Press, New York, 1971.

[13] T. J. Swift and R. E. Cornick, *J. Chem. Phys.*, **37**, 307 (1962).

[14] J. Gordon, *The Organic Chemistry of Electrolyte Solutions*, P9 204, J. Wiley, New York, 1975.

[15] R. M. Fuoss, *J. Am. Chem. Soc.*, **80**, 5059 (1958).

[16] K_0 is given by

$$k_0 = 4 \times \pi \times N \times a^3 \exp[-U(a)/kT]/3000$$

where $U(a)$ is the electrostatic potential at the approach distance a and it can be calculated by applying the Debye-Huckel theory. Here, a = 5 Å has been assumed.

[17] E. Mentasti, E. Pelizzetti, F. Secco and M. Venturini, *Inorg. Chem.*, **18**, 2007 (1979).

[18] T. M. Che and K. Kustin, *Inorg. Chem.*, **20**, 509 (1981).

[19] E. Mentasti, F. Secco and M. Venturini, *Inorg. Chem.*, **19**, 3528 (1980).

[20] L. A. Dominey and K. Kustin, *J. Am. Chem. Soc.*, **23**, 103 (1984).

[21] S. Chopra and R. B. Jordan, *Inorg. Chem.*, **22**, 1708 (1983).

[22] P. Amico, G. Ostacoli and A. Vanni, *Ann. Chim. (Italy)*, **65**, 465 (1975), P. G. Daniele and G. Stacoli, *ibid*, **67**, 465 (1977).

[23] E. Coronel Canaviri, M. R. Feliz and A. L. Capparelli, unpublished results.

[24] T. Inoue, K. Kojima and R. Shimozowa, *Inorg. Chem.*, **22**, 3972 (1983).

[25] T. Inoue, K. Sugahara, K. Kojima and R. Shimozawa, *Inorg. Chem.*, **22**, 3977 (1983).

[26] T. Inoue, K. Kojima and R. Shimazowa, *Bull. Chem. Soc. Jpn.*, **59**, 1683 (1986).

[27] S. Harada, K. Amidaiji, T. Yasunaga, *Bull. Chem. Soc. Jpn.*, **45**, 1752 (1972).

[28] S. Harada, H. Tanabe and T. Yasunaga, *Bull. Chem. Soc. Jpn.*, **46**, 3125 (1973).