Inorg. Chem. **1990**, *29*, 2937-2940 **2937**

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Kinetics and Mechanism of the Linkage Isomerism in the Nickel(II)-2,3-Dihydroxybenzoate System

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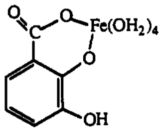
Received May 15, 1989

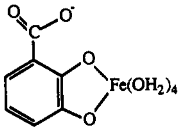
The complexation of nickel(II) by 2,3-dihydroxybenzoate has been studied by stopped-flow spectrophotometry at 25 °C in PIPES buffer between pH 6.5 and 7.0. The apparent pseudo-first-order rate constant appears to decrease with increasing nickel(II) concentration and changes with the observation wavelength between 290 and 350 nm. It is shown that the reaction is biphasic and the kinetic behavior is interpreted in terms of the formation and interconversion of two linkage isomers in which nickel(II) is complexed, as in salicylate or as in catechol.

Introduction

Recent work[1] has shown that aqueous iron(III) complexes with 2,3-dihydroxybenzoate in a salicylate mode (I) in dilute acid ([H^+ > 0.01 M); however, it was possible to estimate that linkage isomerism to the catecholate form (II) might occur at pH >6.



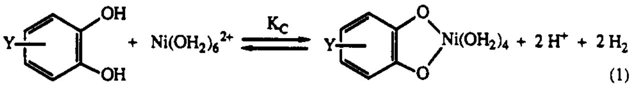
I



II

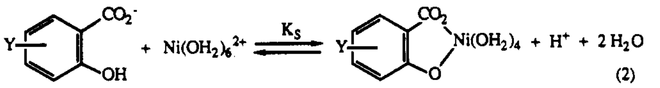
Raymond and co-workers[2] have used deuterium NMR shifts to provide evidence for a similar linkage isomerism on protonation in enterobactin models. These systems differ from the 2,3-dihydroxybenzoate, having a carboxamide rather than carboxylate function.

The kinetics of this process cannot be studied directly in simple aqua-iron(III) systems because of the extensive hydrolysis and possible polymerization of aqueous iron(III) at pH >3. However, aqueous nickel(II) is not seriously affected by these problems for pH <7. On the basis of known formation constants,[3] it is possible to estimate that the equilibrium constant for eq 1 will have log



(1)

$K_C \sim -12.5$ for Y ≡ H, SO_3^-, Cl, and CO_2^-. For salicylate complexes as in eq 2, available data[3,4] indicate that log $K_S \sim -5.7$



(2)

for Y ≡ H and Cl. From the above values, it is possible to estimate the equilibrium constant for the linkage isomerism (K_i) shown in Scheme I. Since $K_i = K_C/K_S$, then log $K_i \sim (-1.25) - (-5.7) = -6.8$. Therefore $K_i \sim 2 \times 10^{-7}$ M, and there could be similar amounts of the catecholate and salicylate complexes present at pH ~ 6.8.

The purpose of this study is to determine if the isomerism in Scheme I does occur, and then to assess the kinetics and reaction sequence. One can imagine that the isomerism might occur by dissociation of one isomer and complexation to form the other or by an intramolecular rearrangement.

Results

Conventional Stopped-Flow Studies. When separate buffered solutions of $Ni(NO_3)_2$ and 2,3-dihydroxybenzoate (DHB) adjusted to the same pH are mixed on a stopped-flow system ([Ni^{2+}] ≫ [DHB], a substantial absorbance change is observed in the 330-nm region. The absorbance-time curves are reasonably represented by a simple exponential function based on visual comparison to simulated exponential curves. However, the apparent pseudo-first-order rate constant decreases with increasing nickel(II) concentration. For example, as the [Ni^{2+}] is changed between 0.02 and 0.05 M, the rate constant changes from 16 to 12 s^{-1} at pH 6.5 and from 13 to 8.6 s^{-1} at pH 6.8 (rate constant uncertainties are ±5-7%, based on reproducibility). Representative observations at pH 6.5 are shown in Figure 1, which clearly illustrates the anomalous concentration dependence of the absorbance-time curves. Such a concentration dependence of the rate law is unprecedented for a complexation reaction on nickel(II) or any other metal ion, as far as we are aware. Furthermore, the apparent rate constant varies with the observation wavelength. When the wavelength is changed in 10-nm increments between 290 and 350 nm, the apparent rate constant decreases monotonically from 16.8 to 9.7 s^{-1} (pH 6.81; [Ni^{2+}], 0.050 M; [DHB] = 4.8×10^{-4} M; in 0.01 M PIPES buffer, 0.3 M $NaNO_3$ at 25 °C).

The above observations provided the first indication that something other than just simple complexing of nickel(II) by DHB must be occurring.

Equilibrium Studies. In an attempt to define the system more completely, potentiometric titrations were carried out on solutions

Scheme I

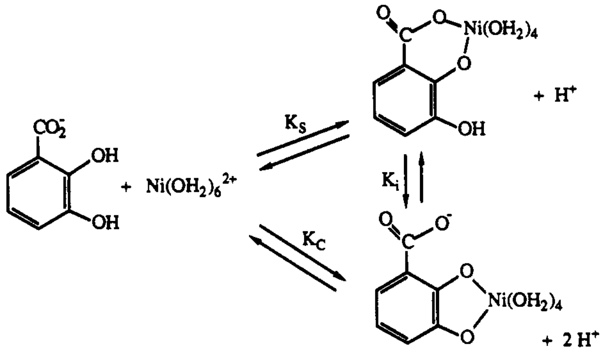


Table I. Rate Constants Obtained on Mixing Aqueous $Ni(NO_3)_2$ and 2,3-Dihydroxybenzoate at 25 °C in 0.30 M $NaNO_3$

		k value at various [Ni^{2+}], s^{-1}				
pH[a]		0.050 M	0.040 M	0.030 M	0.020 M	0.010 M
6.50	k_1(obsd)	18	16	14	13	*b*
	k_1(calcd)	17.4	15.7	14.1	12.4	
	k_2(obsd)[c]	8.1	6.8	4.1	6.8	
6.75	k_1(obsd)	25	20	17	15	13
	k_1(calcd)	23.8	20.9	17.9	15.0	12.0
	k_2(obsd)[c]	10	8.3	8.9	7.6	*d*
6.81	k_1(obsd)	30	22	18	15	11
	k_1(calcd)	26.0	22.6	19.2	15.9	12.5
	k_2(obsd)[c]	7.2	5.8	9.3	7.4	*d*

[a] pH controlled by 0.01 M PIPES buffer in both the nickel(II) and DHB solutions before mixing. [b] The absorbance change was too small to permit an analysis. [c] These values are obtained from fits with k_1 fixed at the appropriate calculated value. [d] The k_2 could not be evaluated because of the small contribution of this term.

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(1) Xu, J.; Jordan, R. B. *Inorg. Chem.* **1988**, *27* 1502.
(2) Cass, M. E.; Garret, T. M.; Raymond, K. N. *J. Am. Chem. Soc.* **1989**, *111*, 1677.
(3) Martell, A. E.; Smith, R. M. *Critical Stability Constants*; Plenum: New York, 1977; Vol. 3.
(4) Chopra, S.; Jordan, R. B. *Inorg. Chem.* **1983**, *22*, 1708.

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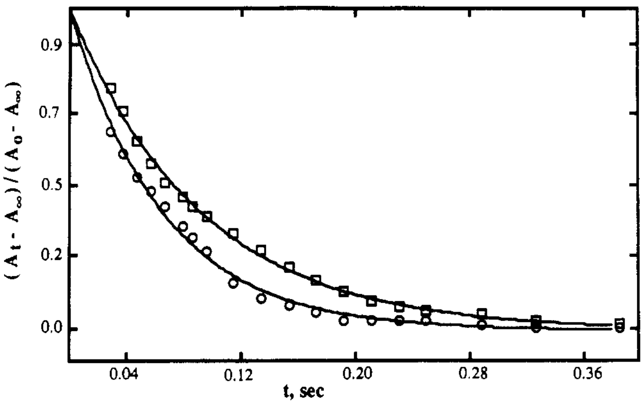
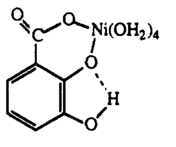
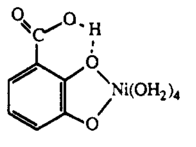


Figure 1. Variation of absorbance change at 330 nm with time for the reaction of $Ni(OH_2)_6^{2+}$ with 2,3-dihydroxybenzoate at pH 6.5 in 0.30 M $NaNO_3$ at 25 °C: (○) 0.020 M Ni^{2+}; (□) 0.050 M Ni^{2+}.

of nickel(II) (0.01–0.05 M) and DHB (9.6 × 10^{-4} M) in 0.30 M $NaNO_3$. These results gave a pK_a for DHB of 2.65 ± 0.07, in good agreement with previous results.[5,6] Since previous work[4] has indicated that a protonated salicylate complex might be present, and a protonated catecholate complex also would seem possible, the titration results have been analyzed by nonlinear least-squares fitting of the titration data in terms of the reactions in Scheme II. In principle, it is possible to determine $K_S + (K_C/K_a'') = K_{2a}\{(k_{12}/k_{21}) + (k_{13}/k_{31})\}$, K_C, and K_S/K_a' from the titrations. In practice, the complexity of the system and the rather weak complexing only allow one to determine the magnitude of $K_S + (K_C/K_a)$ and to establish limits on the other parameters. The analysis, including a correction for the hydrolysis of $Ni(OH_2)_6^{2+}$,[7] leads to the conclusion that $K_S + (K_C/K_a'') \approx 5 \times 10^{-6}$, $K_C < 8 \times 10^{-14}$, and $K_S/K_a' < 1$. Then, the protonated salicylate complex has p$K_a' \sim 5$ and the protonated catecholate complex has p$K_a'' \sim 8$. The former value may be smaller than that of the species with salicylate (6.2) because intramolecular hydrogen bonding stabilizes the anion, as shown in III. On the other hand, hydrogen bonding may favor protonation of the catecholate complex, as shown in IV.


III

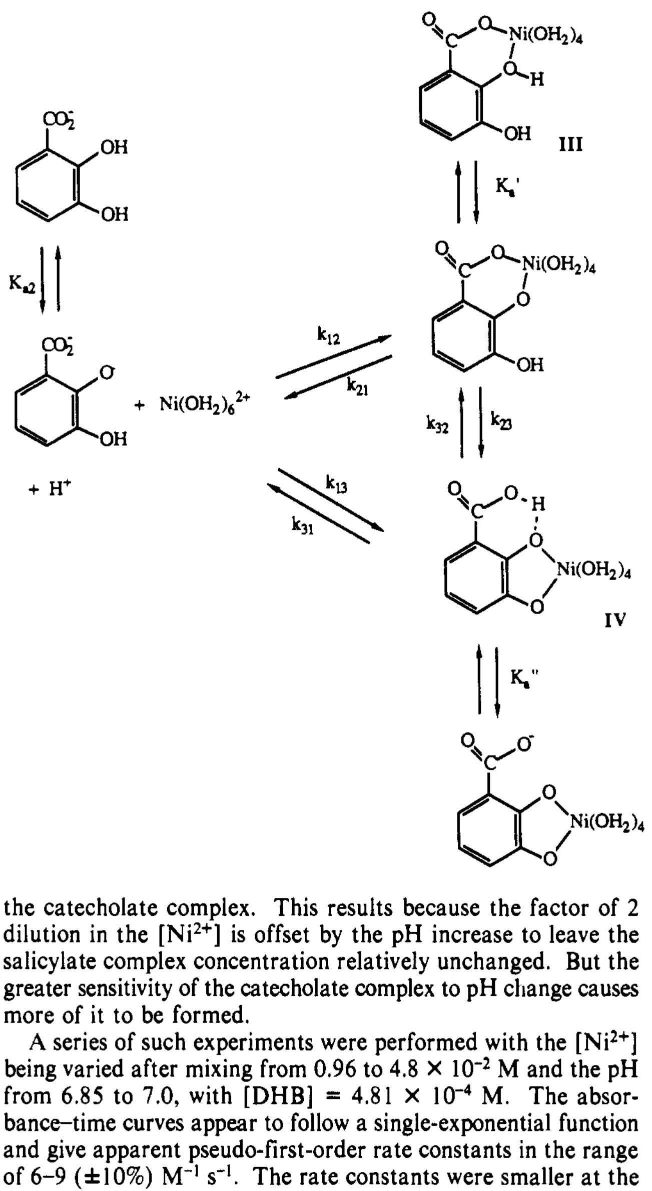

IV

The net result of these titration studies is that they are not able to establish unequivocally the species present although they provide some guidelines to the possible formation constants. This uncertainty is produced by the relative weakness of the complexation and by the potential complexity of the system.

Concentration and pH Jump Studies. The analysis of the conventional stopped-flow observations is complex and ambiguous because there appear to be at least two processes with rather similar rates. It was determined that the description of this system might be put on firmer ground by a combination of concentration and pH changes on mixing in a stopped-flow system.

Since formation of the catecholate complex involves the loss of two protons while the salicylate complex involves one, the catecholate complex is favored at higher pH. Preliminary calculations indicate that, if a solution initially containing some buffer, nickel(II) and DHB ($[Ni^{2+}] \gg [DHB]$) at pH 6.3–6.4 is mixed with an equal volume of solution containing buffer and DHB at the same concentration and pH ~ 7.2, then the final pH can be 6.9–7.0 and the major concentration change will be formation of the catecholate complex. This results because the factor of 2 dilution in the $[Ni^{2+}]$ is offset by the pH increase to leave the salicylate complex concentration relatively unchanged. But the greater sensitivity of the catecholate complex to pH change causes more of it to be formed.

Scheme II



A series of such experiments were performed with the $[Ni^{2+}]$ being varied after mixing from 0.96 to 4.8 × 10^{-2} M and the pH from 6.85 to 7.0, with [DHB] = 4.81 × 10^{-4} M. The absorbance–time curves appear to follow a single-exponential function and give apparent pseudo-first-order rate constants in the range of 6–9 (±10%) M^{-1} s^{-1}. The rate constants were smaller at the lower $[Ni^{2+}]$, where they appear to be almost independent of pH. At higher $[Ni^{2+}]$, the rate constants appear to decrease with increasing pH. These results are given in Table II.

Interpretation and Conclusions

The simplest explanation for a wavelength dependence and unusual dependence on $[Ni^{2+}]$ of the apparent pseudo-first-order rate constant is that more than one process is being observed. If there are two processes, they must make different relative contributions to the observed absorbance, depending on the $[Ni^{2+}]$ and on the wavelength. If one of the processes has the normal first-order dependence on $[Ni^{2+}]$, then its relative contribution must decrease as the $[Ni^{2+}]$ increases. Because the absorbance–time curves appear exponential in this case, it is also implied that the rates for the processes are not greatly different. Under the latter circumstances, it generally is difficult to resolve the independent rate constants from least-squares fits of the absorbance–time data, but such fits were reasonably successful in this case, at least to the extent of indicating the expected behavior semiquantitatively. Our overall goal is to provide an explanation for these observations in terms of reaction pathways and rate constants, which are consistent with the extensive information available[4,8] on the substitution reactions of $Ni(OH_2)_6^{2+}$.

The absorbance–time data from the conventional mixing experiments were fitted to eq 3, which describes a two-process

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(5) Harada, H. *Bull Chem. Soc. Jpn.* **1971**, *44*, 3459.

(6) Avdeef, A.; Sofen, S. R.; Bregrante, J.; Raymond, K. N. *J. Am. Chem. Soc.* **1978**, *100*, 5362.

(7) A hydrolysis constant for $Ni(OH_2)_6^{2+}$ of 6 × 10^{-11} M has been used: Baes, C. F.; Mesmer, R. E. *The Hydrolysis of Cations*; Wiley: New York, 1976; Chapter 10.

(8) Wilkins, R. G. *Acc. Chem. Res.* **1970**, *3*, 408.

Table II. Rate Constants for Two Processes in the Reaction of $Ni(OH_2)_6^{2+}$ + 2,3-Dihydroxybenzoate at 25 °C in 0.30 M $NaNO_3$

		k, s^{-1}	
pH	$10^2[Ni^{2+}]$, M	obsd[a]	calcd
6.5	5.00	18	16.8
6.5	4.00	16	15.4
6.5	3.00	14	14.0
6.5	2.00	13	12.7
6.75	5.00	25	23.1
6.75	4.00	20	20.4
6.75	3.00	17	17.8
6.75	2.00	15	15.2
6.75	1.00	13	13.0
6.81	5.00	30	25.2
6.81	4.00	22	22.0
6.81	3.00	18	19.0
6.81	2.00	15	16.0
6.81	1.00	11	13.4
7.01	0.962	7.21	6.57
6.97	0.962	6.91	6.53
6.93	0.962	6.62	6.49
6.89[b]	0.962	6.99	6.45
6.97	1.92	6.23	7.10
6.93[b]	1.92	6.07	7.05
6.88[b]	1.92	6.60	7.01
7.00	2.88	6.94	7.37
6.93[b]	2.88	6.94	7.32
6.86[b]	2.88	7.04	7.25
7.00	3.85	8.28	7.52
6.95[b]	3.85	8.24	7.50
6.92[b]	3.85	7.83	7.47
6.84	3.85	6.58	7.40
6.99[b]	4.81	8.63	7.61
6.91[b]	4.81	8.92	7.57
6.84	4.81	7.66	7.52

[a] The first 14 entries are γ_+ values given as k_1 in Table I, and the remainder are γ_- (k_2) determined from concentration–pH jump experiments. [b] These results are the average of two independent experiments in which the final pH is within ±0.02 of the value given.

system, where A_t is the absorbance at time t, and A_1, A_2, A_3, k_1, and k_2 are fitting parameters. It was necessary to alternately

$$A_t = A_1 e^{-k_1 t} + A_2 e^{-k_2 t} + A_3 \tag{3}$$

fix k_1 or k_2 and determine the best fit value of the other that gives the minimum standard error of the fit. In general, k_1 is well-defined at low $[Ni^{2+}]$ and k_2 is well-defined at high $[Ni^{2+}]$, and these limits, along with the assumption of a linear dependent of k_1 on $[Ni^{2+}]$, have been useful in narrowing the range of fixed values for the fitting process. Fortunately, these values also produce the minimum standard errors on the fits, but none of the fits are so bad as to be rejected on any statistical grounds.

The picture that emerges from the above analysis is that the reaction associated with the larger rate constant (k_1) increases in rate with increasing $[Ni^{2+}]$ and this term produces a larger percentage of the absorbance change at lower $[Ni^{2+}]$. The smaller rate constant (k_2) appears to be almost independent of $[Ni^{2+}]$, and this is the dominant process observed at higher $[Ni^{2+}]$. The values of k_1 can be fitted to a reasonable rate law, given by eq 4.[9] The observed and calculated values of k_1 are compared in

$$k_1 = (5.25 \pm 1.5) \times 10^{-5}\frac{[Ni^{2+}]}{[H^+]} + 9.1 \pm 2.2 \tag{4}$$

Table I. The values of k_2, also given in Table I, do not show any particular trend with pH or $[Ni^{2+}]$ and have an average value of 7.8 ± 1.2 s^{-1}. The form of eq 4 is the same as that expected for complexation of nickel(II) by the conjugate base of a weak acid,[4] as shown in Scheme III. A comparison of the rate law for Scheme III and eq 4 shows that $k_{43}K_a = 5.3 \times 10^{-5}$ and $k_{43}/K_f = 9.1$, where $K_f = k_{43}/k_{34}$. The ratio of these values gives $K_aK_f = 5.8 \times 10^{-6}$, which corresponds to K_s if salicylate complexing is assumed. The value of log (K_aK_f) = −5.2 is close to that of log K_s noted above for other salicylate systems. The pK_a of 9.84 for DHB[5] can be used to estimate that $k_{43} \sim 3.6 \times 10^5$ M^{-1} s^{-1}. This value is similar to that of 2.9×10^5 M^{-1} s^{-1} for the dianion $H_3COPO_3^{2-}$ [8] and a possible value of 1.3×10^5 M^{-1} s^{-1} for one limiting case of the 3,5-dinitrosalicylate dianion.[4] However Diebler et al.[10] have given a lower value of 3.1×10^4 M^{-1} s^{-1} for the latter system. In general, the concentration dependence and magnitude of k_1 seem consistent with salicylate complexing, as shown in eq 2.

Scheme III

$$HL \underset{}{\overset{K_a}{\rightleftharpoons}} L^- + H^+$$

$$L^- + Ni(OH_2)_6^{2+} \underset{k_{34}}{\overset{k_{43}}{\rightleftharpoons}} L\text{-}Ni(OH_2)_5^+ + H_2O$$

The concentration and pH jump studies were designed to maximize the contribution of the isomerization to the observed absorbance change. Since these studies give rate constants similar to k_2 from the conventional mixing experiments, it is reasonable to associate the k_2 path with the isomerization shown in Scheme I.

These observations provide further encouragement that this system is described by the reactions in Scheme I, and one would like to be able to extract the rate constants for the various steps. A closed solution is available[11] for the kinetic system in Scheme I, and it shows biphasic behavior with the two rate constants ($\gamma_\pm$) given in eq 5, which are mathematically equivalent to k_1 and k_2

$$\gamma_\pm = \frac{S_1 + S_2}{2} \pm \left[\frac{(S_1 - S_2)^2}{4} + (\upsilon_{12} - \upsilon_{32})(\upsilon_{21} - \upsilon_{31})\right]^{1/2} \tag{5}$$

$$S_1 = \upsilon_{12} + \upsilon_{31} + \upsilon_{13}$$

$$S_2 = \upsilon_{21} + \upsilon_{23} + \upsilon_{32}$$

$$S_{12} = k_{12}K_{a2}[Ni]/(K_{a2} + [H^+]),$$
$$\upsilon_{21} = k_{12}K_{a2}K_a'/(K_S(K_a' + [H^+]))$$

$$\upsilon_{13} = k_{13}K_{a2}[Ni]/(K_{a2} + [H^+]),$$
$$\upsilon_{31} = k_{13}K_{a2}A_a''/(K_C(K_a'' + [H^+]))$$

$$\upsilon_{23} = k_{23}K_a'/(K_a' + [H^+]),$$
$$\upsilon_{32} = k_{23}K_S[H^+]/(K_C(K_a'' + [H^+]))$$

in eq 3. The expressions for the reverse rate constants are given in terms of the forward rate constants and the appropriate equilibrium constants to minimize the number of variables. If the reactions in Scheme I properly represent the system, then there must be a set of specific rate constants and equilibrium constants that will reproduce the values of both γ_+ and γ_- at the various experimental conditions of pH and $[Ni^{2+}]$. It may be possible to determine if the intramolecular step (k_{23},k_{32}) is making a significant contribution to the isomerization relative to paths proceeding through free ligand and metal ion.

Although eq 5 appears to predict a complex dependence of the observed rate constants on $[H^+]$ and $[Ni^{2+}]$, further analysis shows that reasonably simple expressions give close approximations of γ_+ and γ_- that are useful in understanding what one can expect from the analysis. A binomial expansion of the square-root term in eq 5 yields the following expressions:

$$\gamma_+ = S_1 + (\upsilon_{12} - \upsilon_{32})(\upsilon_{21} - \upsilon_{31})/(S_1 - S_2) \tag{6}$$

$$\gamma_- = S_2 - (\upsilon_{12} - \upsilon_{32})(\upsilon_{21} - \upsilon_{31})/(S_1 - S_2) \tag{7}$$

Further numerical analysis reveals that, to a reasonable approximation (±15%), the second term on the right-hand side of eqs

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(9) Errors quoted are 95% confidence limits, which are about 3 times larger than 1 standard deviation.

(10) Diebler, H.; Secco, F.; Venturini, M. *J. Phys. Chem.* **1987**, *91*, 5106.

(11) Bamford, C. H.; Tipper, C. F. H. *Chemical Kinetics*; Elsevier: New York, 1969; Vol. 2, pp 31–32.

6 and 7 is small relative to the first term so that

$$\gamma_+ \approx S_1 = \frac{(k_{12} + k_{13})K_{a2}[\mathrm{Ni}^{2+}]}{(K_{a2} + [\mathrm{H}^+])} + \frac{k_{13}K_{a2}[\mathrm{H}^+]}{K_\mathrm{C}(K_a'' + [\mathrm{H}^+])} \quad (8)$$

$$\gamma_- \approx S_2 = \frac{k_{12}K_{a2}K_a'}{K_\mathrm{S}(K_a' + [\mathrm{H}^+])} + k_{23}\left(\frac{K_a'}{(K_a' + [\mathrm{H}^+])} + \frac{K_\mathrm{s}[\mathrm{H}^+]}{K_\mathrm{C}(K_a'' + [\mathrm{H}^+])}\right) \quad (9)$$

Since $K_{a2} = 1.5 \times 10^{-10} \ll [\mathrm{H}^+]$, inspection of eq 8 shows that γ_+ will have a dependence on $[\mathrm{Ni}^{2+}]$ and $[\mathrm{H}^+]$ consistent with eq 4 if $K_a'' \ll [\mathrm{H}^+]$ and K_C is large enough or k_{13} is small enough so that the last term in eq 8 is not dominant. On the other hand, eq 9 predicts that γ_- should be independent of $[\mathrm{Ni}^{2+}]$ because the last term in eq 7 disappears in the first-order approximation used to obtain eq 9. The magnitude of γ_- depends directly on k_{12} and k_{23}, but the former may be established by γ_+ so that at least an upper limit on the value of k_{23} is determined from γ_-. In addition, γ_- will not show much pH dependence if $K_a' > [\mathrm{H}^+]$ and $K_a'' \ll [\mathrm{H}^+]$. It is noteworthy that if K_C is decreased in eq 8, then k_{13} will decrease to maintain the magnitude of the last term and k_{12} will increase to keep the first term constant. But the latter increase may be limited by the magnitude of γ_-. It should be apparent that it will be possible to evaluate $k_{12} + k_{13}$ from eq 8, but the separation of the terms will depend on K_C in eq 8 and K_s in eq 9.

Qualitatively k_1 (γ_+) varies with $[\mathrm{Ni}^{2+}]$ and pH as predicted by eq 8 and k_2 (γ_-) is relatively independent of these variables as required by eq 9. The values of k_1 (γ_+) and k_2 (γ_-) have been fitted simultaneously by a nonlinear least-squares method to eq 5 while various constraints were imposed on the equilibrium constants, as indicated by the pH titration results. The experimental and calculated results are compared in Table II. As expected from the above discussion, the value of $k_{12} + k_{13}$ is well-defined at $(3.5 \pm 1.1) \times 10^5$ M^{-1} s^{-1}. A "best fit" is obtained with $K_\mathrm{s} + (K_\mathrm{C}/K_a'') = 6 \times 10^{-6}$ and $K_\mathrm{s} = 2.5 \times 10^{-6}$, which gives $k_{12} = (1.9 \pm 0.7)$ ′ 10^5 M^{-1} s^{-1} and $k_{13} = (1.4 \pm 0.7) \times 10^5$ M^{-1} s^{-1}, while k_{23} is undefined at 0.28 ± 4.8 s^{-1}. The latter value indicates that the intramolecular rearrangement path is not significant in the linkage isomerism process in this system. The values of k_{12} and k_{13} are similar, as expected for a dissociative ion pair substitution on $\mathrm{Ni(OH_2)_6}^{2+}$, and their magnitude is consistent with that expected for reaction of a dianion.[8]

In summary, it has been found that the complexing of $\mathrm{Ni(OH_2)_6}^{2+}$ by 2,3-dihydroxybenzoate is not a simple kinetic process and that the complexity can be accounted for by the formation of salicylate and catecholate linkage isomers. In this labile system, the linkage isomerism proceeds predominantly by dissociation and recomplexation rather than by an intramolecular rearrangement.

Experimental Section

Materials. All solutions for kinetic and equilibrium studies were prepared in deionized water distilled from alkaline permanganate in an all-glass apparatus. The nickel(II) solutions were prepared by dissolving reagent grade nickel nitrate (Fisher) in water and standardized by EDTA titration.

The ligand solutions were made by dissolving weighed amounts of 2,3-dihydroxybenzoic acid (Aldrich) in water and diluting to volume. These solutions were prepared fresh daily and stored under argon. When required, the solutions contained PIPES buffer (Sigma) and the pH was adjusted to the desired value before dilution to known volume. All solutions contained enough $\mathrm{NaNO_3}$ to give an ionic strength of 0.30 M.

For the titration studies, solutions containing 9.62×10^{-4} M DHB and various concentrations of nickel(II) (0.010–0.050 M) in 0.30 M $\mathrm{NaNO_3}$ were titrated with 0.0102 M NaOH. The pH meter was calibrated to give $[\mathrm{H}^+]$ by titrating 0.0100 M HCl with the same NaOH between pH 2 and 5.5.

Instrumentation. Preliminary spectrophotometric observations were done on Cary 219 and Hewlett-Packard 8451 spectrophotometers. The pH was determined on a Corning 124 pH meter standardized with appropriate buffers. The stopped-flow observations were done on an Aminco-Morrow stopped-flow system. The rate constants have been determined from the average of five to seven runs with the same solutions under a particular set of conditions.

Acknowledgment. We thank the Natural Sciences and Engineering Research Council of Canada for financial support, and L.C. thanks the Coordination Chemistry Institute of Nanjing University for granting a study leave to do this work.

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Characterization of Pentacyanoferrate(II) and -(III) Complexes of Adenosine and Related Aminopyridine Ligands

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Received January 5, 1989

A series of $\mathrm{Fe(CN)_5L^{3-}}$ complexes, where L = adenosine, 1-methyladenosine, tubercidin, and 2- and 3-aminopyridines, were prepared and characterized in aqueous solution. A metal to ligand charge-transfer transition was observed for complexes of adenosine (337 nm), tubercidin (340 nm), 2-aminopyridine (345 nm), and 3-aminopyridine (362 nm). Their corresponding Fe(III) complexes also display a ligand to metal charge-transfer absorption at 570 (adenosine), 635 (tubercidin), 655 (2-aminopyridine) and 695 nm (3-aminopyridine). The 1-methyladenosine complex of pentacyanoferrate(II) exhibits a band maximum at 375 nm, which is similar to that of imidazole complex. The rate constants of formation and dissociation were measured, and the k_f and k_d values (25 °C, $\mu = 0.10$ M ($\mathrm{LiClO_4}$), pH = 8) are 269 M^{-1} s^{-1} and 9.70×10^{-2} s^{-1} (adenosine), 257 M^{-1} s^{-1} and 0.612 s^{-1} (1-methyladenosine), 70.4 M^{-1} s^{-1} and 0.596 s^{-1} (tubercidin), 401 M^{-1} s^{-1} and 0.585 s^{-1} (2-aminopyridine), and 310 M^{-1} s^{-1} and 1.84×10^{-3} s^{-1} (3-aminopyridine). Cyclic voltammetry of the complexes under study has shown that the oxidation is a one-electron reversible process with $E_{1/2}$ values of 0.49 (adenosine), 0.44 (tubercidin), 0.41 (2-aminopyridine), and 0.43 V (3-aminopyridine) vs. NHE at 25 °C, $\mu = 0.10$ M ($\mathrm{LiClO_4}$), and pH = 8 (Tris). Both spectral and electrochemical results suggest that the coordination site of both $\mathrm{Fe(CN)_5ado^{3-}}$ and $\mathrm{Fe(CN)_5ado^{2-}}$ complexes is at N-1 of the nucleic acid and there is no linkage isomerization from N-1 to the exocyclic nitrogen when the metal center is oxidized from Fe(II) to Fe(III).

The complexes of pentaammineruthenium(II) and -(III) with various nucleic acids have been well characterized during the past decade.[1-5] One of the advantages of using the ruthenium ammine metal center is its substitution inertia in both the 2+ and 3+ oxidation states, which may simplify the experimental results and provide a way to systematically study the effect of the metal ion

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0020-1669/90/1329-2940$02.50/0