Kinetics of Dissociation of Iron(III) Complexes of Tiron in Aqueous Acid

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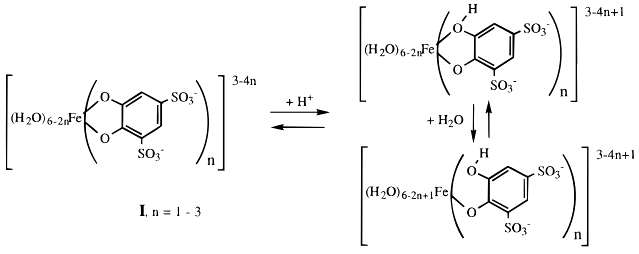
The kinetics of dissociation of the mono, bis, and tris complexes of Tiron (1,2-dihydroxy-3,5-benzenedisulfonate) have been studied in acidic aqueous solutions in 1.0 M $HClO_4$/$NaClO_4$, as a function of $[H^+]$ and temperature. In general, the kinetics can be explained by two reactions, $(H_2O)Fe(L)_{n-1} + H_2L \rightleftarrows (H_2O)Fe(L_nH) + H^+$ (k_n, k_{-n}) and $(HO)Fe(L)_{n-1} + H_2L \rightleftarrows (H_2O)Fe(L_nH)$ (k_n', k_{-n}'), a rapid equilibrium, $(H_2O)Fe(L_nH) \rightleftarrows (H_2O)Fe(L)_n + H^+$ (K_{cn}), and the formation constant $(H_2O)Fe(L)_{n-1} + H_2L \rightleftarrows (H_2O)Fe(L)_n + 2H^+$. For $n = 1$, the reaction was observed at 670 nm, and at $[H^+]$ of 0.05−0.5 M at temperatures of 2.0, 14.0, 25.0, and 36.7 °C. For $n = 2$, the analogous conditions are 562 nm, at $[H^+]$ of 1.5×10^{-3} to 1.4×10^{-2} M at temperatures of 2.0, 9.0, and 14.0 °C. For $n = 3$, the conditions are 482 nm, at pH 4.5−5.7 in 0.02 M acetate buffer at temperatures of 1.8, 8.0, and 14.5 °C. The rate or equilibrium constants (25 °C) with ΔH^* or $\Delta H°$ (kcal mol^{-1}) and ΔS^* or $\Delta S°$ (cal mol^{-1} K^{-1}) in brackets are as follows: for $n = 1$, $k_1 = 2.3$ M^{-1} s^{-1} (8.9, −27.1), $k_{-1} = 1.18$ M^{-1} s^{-1} (4.04, −44.8), $K_{c1} = 0.96$ M (−9.99, −33.6), $K_{f1} = 2.01$ M (−5.14, −15.85); for $n = 2$, $k_{-2}/K_{c2} = 1.9 \times 10^7$ (19.9, 41.5) and $k_{-2}'/K_{c2} = 1.85 \times 10^3$ (1.4, −38.8) and a lower limit of $K_{c2} > 0.015$ M; for $n = 3$, $k_3 = 7.7 \times 10^3$ (15.8, 12.3), $k_{-3} = 1.7 \times 10^7$ (16.2, 28.9), $K_{c3} = 7.4 \times 10^{-5}$ M (4.1, −5.1), and $K_{f3} = 3.35 \times 10^{-8}$ (3.7, −21.7). From the variations in rate constants and activation parameters, it is suggested that the $Fe(L)_2$ and $Fe(L)_3$ complexes undergo substitution by dissociative activation, promoted by the catecholate ligands.

Introduction

There have been numerous studies[1] on the complexation of aqueous iron(III), and the general conclusion is that $Fe(OH_2)_6{}^{3+}$ is much less reactive than its conjugate base, $(H_2O)_5Fe(OH)^{2+}$, and that the latter reacts by dissociative activation, while the former shows associative characteristics. There are far few studies on the substitution reactivity of other iron(III) complexes, and therefore, the reactivity and mechanistic effects of other ligands on the iron(III) center are not well characterized. The main exception to this generalization is the hydroxamic acid complexes that are models for desferrioxamine B. These have been the subject of several studies by Crumblis and co-workers[2,3] and Birus and co-workers.[4,5] The hydrolysis of the tris-, bis-, and monoacetohydroxamate complexes has been kinetically characterized by Birus et al.,[5] and the pressure dependence has been reported recently.[4]

Iron(III) catechol complexes have been widely investigated as models for the siderophile enterobactin by Raymond and co-workers[6] and as dioxygenase models by Que and co-workers.[7] The simple catechol systems are less amenable to systematic hydrolysis studies because of redox instability involving iron(III) or iron(III) and dioxygen. The early work of McBryde[8] indicates that oxidation should not be a problem with Tiron (1,2-dihydroxy-3,5-benzenedisulfonate), that the mono, bis, and tris complexes (**I**) are readily distinguished spectrophotometrically.



I, n = 1 - 3

II

In addition, the formation constants (see species distribution diagram in supporting information) should allow for independent study of these complexes by an appropriate choice of starting pH and Tiron concentration. The present study was undertaken to determine the reactivity patterns of these catecholate complexes.

Since protonated catecholate complexes of iron(III) have been proposed as critical intermediates in dioxygenase models,[7,9] it was also of interest to determine if species such as **II** might be characterized by the kinetic analysis. Que and co-workers[10] first structurally characterized such a monodentate complex, and another example has been reported recently by Kitajima et al.[11]

Results

Preliminary experiments confirmed McBryde's earlier observations[8] that the tris-, bis-, and mono-Tiron complexes of iron(III) are readily distinguished spectrophotometrically. This is shown, along with the time evolution, in Figure 1. The wavelength of the absorbance maxima increases progressively

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(1) Birus, M.; Kujundzic, N.; Probanic, M. *Prog. React. Kinet.* **1993**, *18*, 171.

(2) Crumblis, A. L. *Coord. Chem. Rev.* **1990**, *105*, 155. Batinic-Haberle, I.; Spasojevic, I.; Crumblis, A. L. *Inorg. Chem.* **1994**, *33*, 3151.

(3) Caudle, M. T.; Crumblis, A. L. *Inorg. Chem.* **1994**, *33*, 4077 and references therein.

(4) Birus, M.; van Eldik, R. *Inorg. Chem.* **1991**, *30*, 4559 and references therein.

(5) Birus, M.; Bradic, Z.; Kujundzic, N.; Pribanic, M.; Wilkins, P. C.; Wilkins, R. G. *Inorg. Chem.* **1985**, *24*, 3980.

(6) Karpishin, T. B.; Stack, T. D. P.; Raymond, K. N. *J. Am. Chem. Soc.* **1993**, *15*, 182, 6115. Borgias, B. A.; Barclay, S. J.; Raymond, K. N. *J. Coord. Chem.* **1986**, *15*, 109.

(7) Que, L., Jr.; Kolanczyk, R. C.; White, L. S. *J. Am. Chem. Soc.* **1987**, *109*, 5373 and references therein.

(8) McBryde, W. A. E. *Can. J. Chem.* **1964**, *42*, 1917.

(9) Lauffer, R. B.; Heistand, R. H.; Que, L., Jr. *J. Am. Chem. Soc.* **1981**, *103*, 3947.

(10) Heistand, R. H.; Roe, H. A. L.; Que, L., Jr. *Inorg. Chem.* **1982**, *21*, 676.

(11) Kitajima, N.; Ito, M.; Fukui, H.; Moro-oka, Y. *J. Am. Chem. Soc.* **1993**, *115*, 9335.

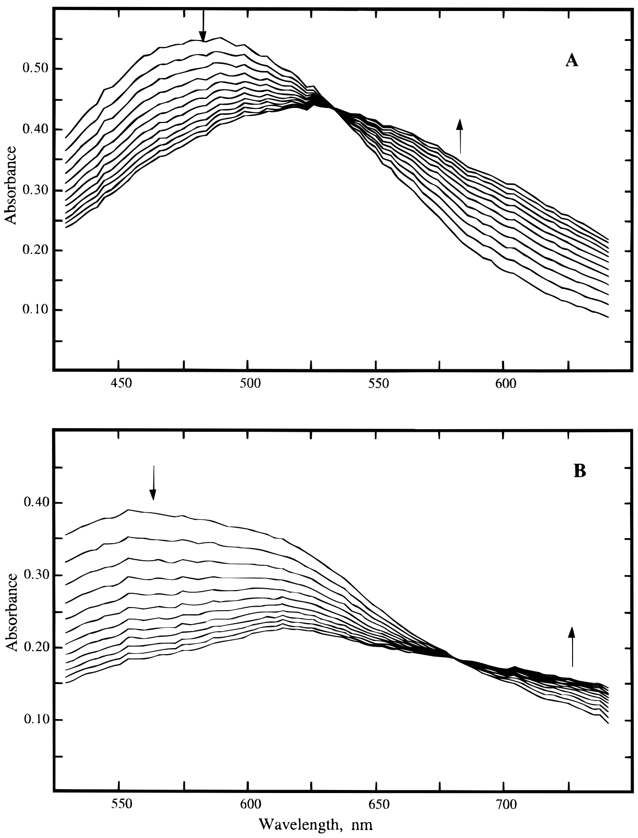


Figure 1. Variation of absorbance with time and wavelength for the hydrolysis of Tiron complexes of aqueous iron(III) in 1.0 M $NaClO_4$/$HClO_4$ at 14.0 °C and 0.01 s time intervals. A: conversion of tris to bis at pH 4.8 in 0.02 M acetate buffer at final total concentrations of iron(III) = 9.6×10^{-5} M and Tiron = 5.8×10^{-4} M. B: conversion of bis to mono at $[H^+] = 1.53 \times 10^{-3}$ M and at final total concentrations of iron(III) = 9.6×10^{-5} M and Tiron = 5.8×10^{-4} M.

from the tris to bis to mono complexes, and the excellent isosbestic behavior for the tris−bis and bis−mono conversions also is shown in Figure 1. After analyzing the kinetic results for all three reactions, it appears that they can be described in general by Scheme 1, where H_2L represents the Tiron dianion protonated on the OH groups, only one of the possible water ligands is included, and charges are omitted for generality.

The reaction described by K_{cn} might be proton dissociation from a protonated chelate or from a monodentate form of the catechol derivative (**II**) to give the chelate with proton loss. Therefore the process may involve both deprotonation and chelation. The predicted first-order rate constant for Scheme 1 can be expressed in terms of either the complex formation rate constants (k_n) or the dissociation constants (k_{-n}), as given by eqs 1 and 2, respectively,

$$k_{obsd} = (k_n[H^+] + k_n'K_{m(n-1)})\left\{\frac{[H_2L]}{K_{m(n-1)} + [H^+]} + \frac{K_{cn}[H^+]}{K_{fn}(K_{cn} + [H^+])}\right\} \quad (1)$$

$$k_{obsd} = (k_{-n}[H^+] + k_{-n}')\left\{\left(\frac{K_{fn}}{K_{cn}}\right)\frac{[H_2L]}{K_{m(n-1)} + [H^+]} + \frac{[H^+]}{(K_{cn} + [H^+])}\right\} \quad (2)$$

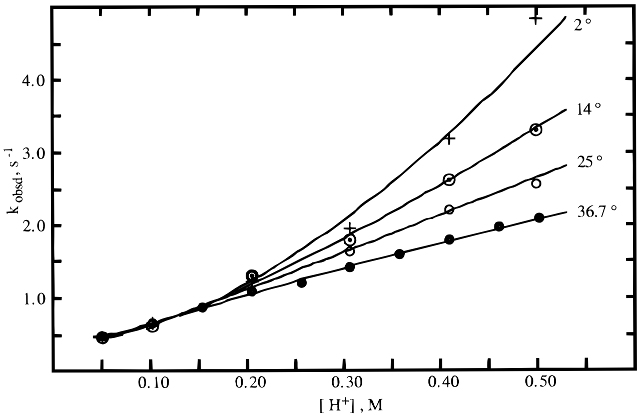
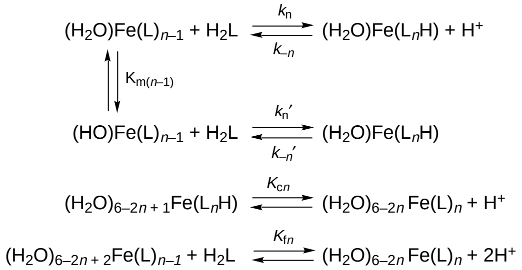


Figure 2. Variation of the first-order rate constant for dissociation of the mono-Tiron complex of iron(III) with $[H^+]$ at 2.0, 14.0, 25.0, and 36.7 °C. Values have been scaled to the same magnitude by multiplying by factors of 85, 17, and 4.2 at 2.0, 14.0, and 25.0 °C, respectively. The total iron(III) and Tiron concentrations are 1.0×10^{-4} and 8.0−9.0 $\times 10^{-4}$ M, respectively, and the reaction was studied at 670 nm.

Scheme 1



$$(H_2O)Fe(L)_{n-1} + H_2L \xrightleftharpoons[k_{-n}]{k_n} (H_2O)Fe(L_nH) + H^+$$

$$\Big\updownarrow K_{m(n-1)}$$

$$(HO)Fe(L)_{n-1} + H_2L \xrightleftharpoons[k_{-n}']{k_n'} (H_2O)Fe(L_nH)$$

$$(H_2O)_{6-2n+1}Fe(L_nH) \xrightleftharpoons{K_{cn}} (H_2O)_{6-2n}Fe(L)_n + H^+$$

$$(H_2O)_{6-2n+2}Fe(L)_{n-1} + H_2L \xrightleftharpoons{K_{fn}} (H_2O)_{6-2n}Fe(L)_n + 2H^+$$

where $[H_2L]$ represents total uncomplexed Tiron which is in the protonated form at the acidities of most of this study. These equations assume that $[H_2L]$ is constant during the reaction, and analysis adjustments for this feature are described in the Experimental Section.

The forward and reverse rate constants are related to the equilibrium constants by eq 3. The predicted k_{obsd} has various

$$\frac{k_n}{k_{-n}} = \frac{K_{fn}}{K_{cn}} = \frac{k_n'}{k_{-n}'}K_{m(n-1)} \quad (3)$$

limiting forms depending on the magnitudes of K_{cn} and $K_{m(n-1)}$ relative to $[H^+]$ and the magnitudes of K_{cn} and K_{fn}. The forms required to fit the dissociation kinetics of the mono, bis, and tris complexes will be discussed individually.

Dissociation Kinetics of (H_2O)Fe(L). This reaction was followed at 670 nm, in the region of the broad absorbance maximum for the complex, by mixing a solution of Tiron ((1.6−1.8) $\times 10^{-3}$ M) and $Fe(ClO_4)_3$ (2×10^{-4} M) at pH 2.7−3 with an equal volume of a solution of $HClO_4$/$NaClO_4$ at concentrations to give the required final acidity and 1.00 M ionic strength. The $[H^+]$ range was 0.05−0.5 M at temperatures of 2.0, 14.0, 25.0, and 36.7 °C, and the experimental first-order rate constants, k_{obsd}, are given in Table S1 of the supporting information.

The variation of k_{obsd} with $[H^+]$ in terms of Scheme 1 at the four temperatures proved to be more complex than would appear from any one temperature. The problem is illustrated by the plots in Figure 2. At 36.7 °C, the k_{obsd} versus $[H^+]$ plot shows a slight leveling at higher $[H^+]$; at 25 °C the plot is essentially linear but shows increasing upward curvature at 14 and 2 °C. In eq 1 or 2, the second term in braces dominates, and the leveling at higher $[H^+]$ and 36.7 °C would indicate that $[H^+] \geq K_{c1}$ and that $k_{-1}' > k_{-1}[H^+]$. However, the upward curvature at the lower temperatures implies that $[H^+] \leq K_{c1}$ and that $k_{-1}[H^+]$ is contributing significantly. With the assumption that a uniform rate law applies at all temperatures, this behavior can only be accounted for by the temperature variation of the

Table 1. Summary of Results for the Reactions of the Iron(III)−Tiron System in 1.0 M $HClO_4/NaClO_4$

parameter	value (25°C)	ΔH,[a] kcal mol^{-1}	ΔS,[a] cal mol^{-1} K^{-1}
$k_1{}^b$ (M^{-1} s^{-1})	2.3	8.89	−27.05
k_{-1} (M^{-1} s^{-1})	1.18	$4.0_4 \pm 1$	-44.8 ± 3
$k_1{}'^{\,b}$ (M^{-1} s^{-1})	1.7×10^3	10.1	−9.9
k_{-1}' (s^{-1})	1.28	14.2 ± 0.2	-10.4 ± 1
K_{c1} (M)	0.96	$-9.9_9 \pm 1$	-33.6 ± 3
K_{f1} (M)	2.01	$-5.1_4 \pm 1$	$-15.8_5 \pm 3.4$
$k_2{}^b$ (M^{-1} s^{-1})	3.2×10^3	19.2	21.9
k_{-2}/K_{c2} (M^{-2} s^{-1})	1.88×10^7	$19.8_9 \pm 0.3$	$41.4_6 \pm 1.6$
$k_{-2}{}^c$ (M^{-1} s^{-1})	$>2.8 \times 10^5$		
k_{-2}'/K_{c2} (M^{-1} s^{-1})	1.85×10^3	$1.4_1 \pm 6$	$-38.8_4 \pm 18$
$k_{-2}{}'^{\,c}$ (s^{-1})	$>2.8 \times 10^1$		
K_{c2} (M)	>0.015		
$K_{f2}{}^d$ (M)	1.7×10^{-4}	(−0.7)	(−19.59)
k_3 (M^{-1} s^{-1})	7.7×10^3	$15.8_3 \pm 0.6$	12.3 ± 3.2
k_{-3} (M^{-1} s^{-1})	1.7×10^7	16.2 ± 0.16	$28.9 \pm .075$
K_{c3} (M)	7.4×10^{-5}	4.1 ± 1	-5.1 ± 4
K_{f3} (M)	3.35×10^{-8}	3.7	−21.7

[a] These are activation parameters ΔH^* and ΔS^* for rate constants and $\Delta H°$ and $\Delta S°$ for equilibrium constants. [b] Calculated from $k_n = k_{-n}K_{fn}/K_{cn}$ and $k_{-n}'K_{fn}/(K_{cn}K_{m(n-1)})$. [c] Lower limits estimated from the lower limit on K_{c2}. [d] $\Delta H°$ estimated from values for K_{f1} and K_{f3} and using the K_{f2} reported by McBryde at 25°C to obtain $\Delta S°$.

various parameters being such that different approximate limiting conditions apply at the temperature extremes. Therefore, the k_{obsd} values at all temperatures were fitted by assuming that the temperature dependence of the rate constants is described by the transition state equation, and the equilibrium constants are described by the usual thermodynamic relationship. The temperature dependence of K_{m0} is known independently and was taken from values given by Baes and Messmer.[12] The resulting fit is very good, and the experimental and calculated values of k_{obsd} are given in Table S1 of the supporting information.

The best fit values of ΔH^* and ΔS^* and $\Delta H°$ and $\Delta S°$ are given in Table 1. At 25 °C, the calculated values are as follows: $k_{-1} = 1.18$ M^{-1} s^{-1}, $k_{-1}' = 1.28$ s^{-1}, $K_{f1} = 2.0$ M, and $K_{c1} = 0.96$ M. These can be combined in eq 3 to calculate the formation rate constants at 25 °C as $k_1 = 2.3$ M^{-1} s^{-1} and $k_1' = 1.7 \times 10^3$ M^{-1} s^{-1} (using $K_{m0} = 1.6 \times 10^{-3}$ M). Similar combination can be used to obtain the ΔH^* and ΔS^* for k_1 and k_1' given in Table 1. The value of K_{f1} is in good agreement with that of 1.8 M determined by McBryde,[8] and k_1' agrees with the values of 2.1×10^3 and 2.2×10^3 M^{-1} s^{-1} determined from formation kinetics by Mentasti et al.[13] and Chatlas and Jordan,[14] respectively. McBryde[8] suggested values of K_{c1} between 0.1 and 0.015 M, depending on assumptions used in the analysis of spectrophotometric data. If such small values are used in eq 3, then the calculated k_1' is 10−70 times larger than that from previous work.[13,14]

Dissociation Kinetics of $(H_2O)Fe(L)_2$. This reaction was studied at 562 nm by mixing solutions of $Fe(ClO_4)_3$ (1.0×10^{-4} M) and Tiron (2.6×10^{-4} M) at pH 4.3 with an equal volume of appropriate solutions of $HClO_4/NaClO_4$. Measurable rates were obtained for $[H^+]$ between 1.5×10^{-3} and 1.4×10^{-2} M at 2.0, 9.0, and 14.0 °C. The acidities are lower for the bis complex compared to those of the mono complex because of the faster hydrolysis of the bis complex. The experimental k_{obsd} values are given in Table S2 of the supporting information and are plotted in Figure 3.

From Figure 3A, it is apparent that k_{obsd} has essentially a second-order dependence on $[H^+]$ at all three temperatures. However, the k_{obsd} at low $[H^+]$ are underpredicted by a simple $[H^+]^2$ dependence, and this is shown by the k_{obsd} versus $[H^+]^{-2}$ plot in Figure 3B. This should give a horizontal line for a purely $[H^+]^2$ dependence, but the plots have a persistent positive slope with a slight temperature dependence. This slope indicates that there is another term in the rate law which becomes significant at low $[H^+]$, but the difficult feature turns out to be the small temperature dependence of this term.

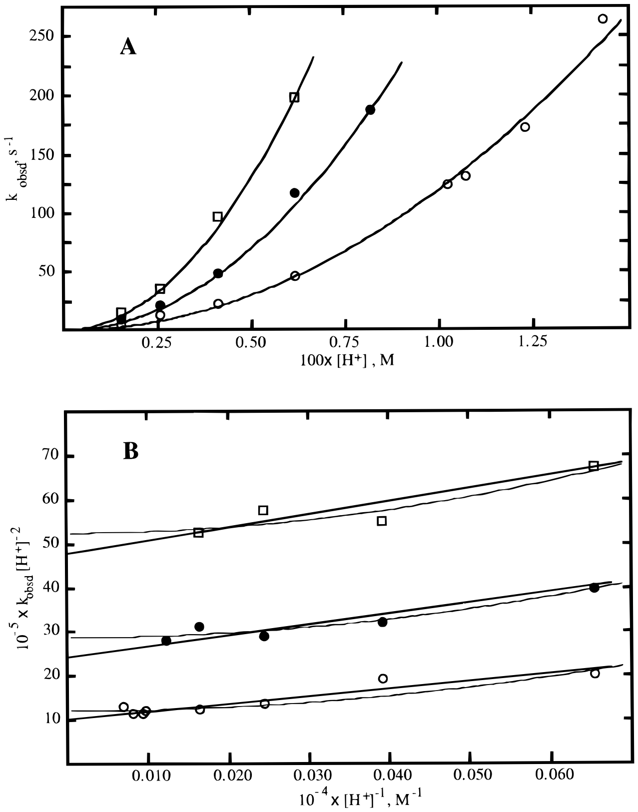


Figure 3. Variation of the first-order rate constant for dissociation of the bis-Tiron complex of iron(III): A with $[H^+]$ at 2.0, 9.0, and 14.0 °C; B variation of $k_{obsd}[H^+]^{-2}$ with $[H^+]^{-1}$. The lines and curves are the results of least-squares fitting to eqs 4 and 5, respectively. The total iron(III) and Tiron concentrations are 5.0×10^{-5} and 1.3×10^{-4} M, respectively, and the reaction was studied at 562 nm.

There are three models that can explain the $[H^+]$ dependence of k_{obsd}. In terms of Scheme 1, if the second term in braces in eq 1 or 2 is dominant, and $K_{c2} \gg [H^+]$ (because the $[H^+]^2$ dependence shows no attenuation at the highest $[H^+]$ values), then eq 2 reduces to eq 4 from which only values of k_{-2}/K_{c2} ($=k_2/K_{f2}$) and k_{-2}'/K_{c2} ($=k_2'K_{m1}/K_{f2}$) can be determined. Least-

$$k_{obsd} = \left(\frac{k_{-2}[H^+]}{K_{c2}} + \frac{k_{-2}'}{K_{c2}}\right)[H^+] \tag{4}$$

squares analysis of all the k_{obsd} values at all three temperatures gives apparent activation parameters ΔH^* (kcal mol^{-1}) and ΔS^* (cal mol^{-1} K^{-1}) of 19.8 ± 0.37 and 41.2 ± 1.8 for k_{-2}/K_{c2} and 4.2 ± 4.7 and -28.5 ± 16 for k_{-2}'/K_{c2}. The parameters for k_{-2}'/K_{c2} are poorly defined because this term makes a small contribution, as can be seen from the small slopes of the lines in Figure 3B. At 25 °C, the predicted values are $k_{-2}/K_{c2} = 1.9 \times 10^7$ and $k_{-2}'/K_{c2} = 3 \times 10^3$.

There is another form of eq 2 which gives an equally good fit of the data. If the k_{-2}' contribution is negligible, and $K_{c2} \gg [H^+]$ and $K_{m2} \ll [H^+]$, then eq 2 simplifies to eq 5. Least-

$$k_{obsd} = \frac{k_{-2}[H^+]}{K_{c2}}\left(\frac{K_{f2}[H_2L]}{[H^+]} + [H^+]\right) \tag{5}$$

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(12) Baes, C. F.; Messmer, R. E. *The Hydrolysis of Cations*; Wiley: New York, 1976; Chapter 10.5.

(13) Mentasti, E.; Pelizzetti, E.; Saini, G. *J. Inorg. Nucl. Chem.* **1976**, *38*, 785.

(14) Chatlas, J.; Jordan, R. B. *Inorg. Chem.* **1994**, *33*, 3817.

squares analysis (see curves in Figure 3B) gives apparent ΔH^* (kcal mol^{-1}) and ΔS^* (cal mol^{-1} K^{-1}) of 18.87 ± 03 and 38.0 ± 1.3 for k_{-2}/K_{c2} and $\Delta H°$ and $\Delta S°$ of -15.5 ± 2.6 and -68.6 ± 8 for K_{f2}. The predicted values at 25 °C are as follows: $k_{-2}/K_{c2} = 1.8 \times 10^7$ M^{-2} s^{-1} and $K_{f2} = 2.4 \times 10^{-4}$ M. The value of K_{f2} is in good agreement with that of 1.7×10^{-4} M determined by McBryde.[8] Despite this agreement, there is an anomaly with this model because the magnitude of $\Delta H°$ is so large. Such values are more typically in the ±5 kcal mol^{-1} range, as found above for K_{f1} and below for K_{f3}.

Since the K_{f2} determined from eq 5 is of the correct magnitude, it appears that the term containing this factor in eq 2 is of some significance. However, it is unrealistic to expect to extract the temperature dependence of k_{-2}/K_{c2}, k_{-2}'/K_{c2}, and K_{f2} from the data. Therefore, the data have been fitted to the simplified version of eq 2 (assuming $K_{c2} \gg [H^+]$ and $K_{m1} \ll [H^+]$) by taking[8] $K_{f2} = 1.7 \times 10^{-4}$ M at 25 °C with $\Delta H° = -0.7$ kcal mol^{-1}, intermediate between the values for K_{f1} and K_{f3}. Least-squares analysis gives the apparent ΔH^* and ΔS^* for k_{-2}/K_{c2} and k_{-2}'/K_{c2}, summarized in Table 1. As expected, the parameters for k_{-2}'/K_{c2} are poorly defined. The experimental and calculated values of k_{obsd} are given in Table S2 of the supporting information. The predicted values at 25 °C are 1.88×10^7 and 1.85×10^3 for k_{-2}/K_{c2} and k_{-2}/K_{c2}, respectively.

Dissociation Kinetics of Fe(L)$_3$. This reaction was studied by mixing a solution of $Fe(ClO_4)_3$ (1.92×10^{-4} M) and Tiron (1.16×10^{-3} M) at pH 6.64 with an equal volume of a solution of 0.04 M acetate buffer adjusted to the desired pH (4.5−5.7), with both solutions in 1.0 M $NaClO_4$. The disappearance of the red tris complex was monitored at 482 nm, but the same rate constant was obtained by observing the appearance of the blue bis complex at 562 nm. Variation of the total acetate concentration between 0.01 and 0.1 M had no effect on the kinetic observations. The results are given in Table S3 of the supporting information.

It should be noted that the analysis of this reaction has been slightly modified. Firstly, to take into account the first ionization of Tiron (pK_{a1} 7.17), the total Tiron concentration is $[H_2L]_t = [H_2L] + [HL]$ and then $[H_2L] = [H^+][H_2L]_t(K_{a1} + [H^+])^{-1}$. Secondly, the analysis indicates that there is no contribution from a hydrolyzed species (i.e., $K_{m2} \ll [H^+]$ in Scheme 1).

The results can be fitted by Scheme 1 with the modifications noted above for Tiron ionization and omission of coordinated water ionization. The analysis also shows that the k_3' path is not required to fit the results. Then eq 1 can be simplified and modified by the substitution $k_3K_{c3}/K_{f3} = k_{-3}$ to obtain eq 6.

$$k_{obsd} = \frac{k_3[H^+][H_2L]_t}{K_{a1} + [H^+]} + \frac{k_{-3}[H^+]^2}{K_{c3} + [H^+]} \qquad (6)$$

The results were analyzed by least-squares fitting to eq 6 with k_3, k_{-3}, and K_{c3} as parameters at each of the temperatures of 14.5, 8.0, and 1.8 °C and found to give good fits (as shown in Table S3 of the supporting information and in Figure 4) with sensible temperature dependencies for each parameter. The best fits for ΔH^* and ΔS^* and $\Delta H°$ and $\Delta S°$ are given in Table 1. The predicted values at 25 °C are as follows: $k_3 = 7.7 \times 10^3$ M^{-1} s^{-1}, $k_{-3} = 1.7 \times 10^7$ M^{-1} s^{-1}, and $K_{c3} = 7.4 \times 10^{-5}$ M. From these one can calculate that $K_{f3} = 3.35 \times 10^{-8}$ M at 25 °C, in reasonable agreement with 2.7×10^{-8} M determined by McBryde,[8] and the ΔH values and K_{f3} at 25 °C can be combined to give the $\Delta H°$ and $\Delta S°$ for K_{f3} in Table 1.

Discussion

A major purpose of this study is to provide some indication of general reactivity patterns for mono, bis, and tris complexes of catechol ligands with aqueous iron(III). The results are collected in Table 1. As already noted, the equilibrium constant values (K_{fn}) from the kinetics are in good agreement with previously determined values.[8] There also is good agreement with earlier studies[13,14] on the rate constant for formation of the mono(iron(III)−Tiron) complex from $Fe(OH)^{2+} + H_2L$. The present work has added the rate constant for $Fe(OH_2)^{3+} + H_2L$ and the activation parameters for both reactions.

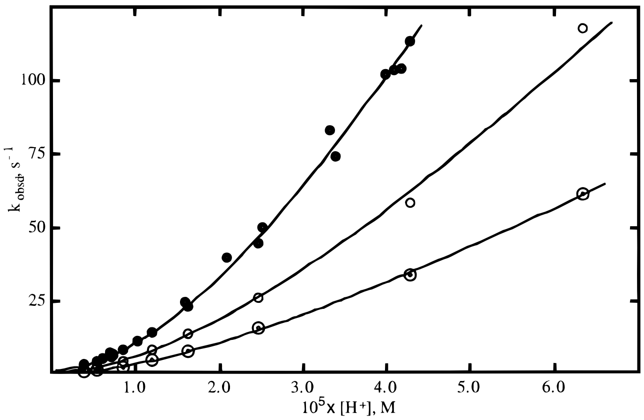


Figure 4. Variation of the first-order rate constant for dissociation of the tris-Tiron complex of iron(III) with $[H^+]$ at 1.8, 8.0, and 14.5 °C. The total iron(III) concentration is 9.6×10^{-5} M, and the Tiron concentration is 5.8×10^{-4} M for most runs, but Tiron concentrations of 9.37×10^{-4}, 7.50×10^{-4}, and 3.75×10^{-4} M also were used. The reaction was studied at 482 nm.

One noteworthy feature is that the k_n' pathway for the hydrolyzed species $(HO)Fe(L)_{n-1}$ is important only for $n = 1$. It is well-established that substitution on aqueous iron(III) is dominated by the reactivity of $(H_2O)_5Fe(OH)^{2+}$, but this pathway is barely detectable for (HO)Fe(L) and of no influence for $(HO)Fe(L)_2$. This may occur because complexation by the catecholate reduces the charge on the iron(III) and thereby reduces the hydrolysis constant ($K_{m(n-1)}$) to an extent that the hydrolyzed species is no longer present at kinetically significant concentrations. Another factor is that the k_n pathway dominates for $n > 1$ because k_2 and k_3 are $>10^3$ times larger than k_1. It appears that complexation by one or two catecholate ligands provides enough labilization so that the effect of an OH^- ligand is attenuated.

The large difference between k_1 and k_2 or k_3 appears to reflect a mechanistic change. For k_1, the ΔS^* is very negative, consistent with the commonly assumed associative activation for substitution of $Fe(OH_2)_6^{3+}$, but the ΔS^* values for k_2 and k_3 are positive, and this overcomes the higher ΔH^* to make these rate constants much larger than k_1. It seems reasonable that the catecholate ligand(s) would favor dissociative activation for k_2 and k_3 by lowering the effective charge on the iron(III) and favoring dissociation of a water ligand. This factor also would disfavor associative nucleophilic attack on the complexed iron(III).

The reverse reactions must have the same mechanistic assignment as the forward reaction. The very negative ΔS^* value for k_{-1} and positive value for k_{-3} are consistent with associative and dissociative activation, respectively, and consistent with the assignments above for k_1 and k_3. Although k_{-2} is not independently defined, the large positive ΔS^* for k_{-2}/K_{c2} suggests that the ΔS^* for k_{-2} is positive and that this pathway is dissociatively activated. The dissociation rate constants also show a large ($>10^5$) increase between k_{-1} and k_{-2}, indicative of a mechanistic change.

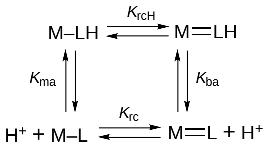
The rate constants for Tiron and two hydroxamic acids reacting with aqueous iron(III) are summarized in Table 2. The values of k_1', k_1, k_2, and k_3 are of similar magnitude for Tiron

Table 2. Comparison of Rate Constants for Tiron and Hydroxamate Complexes of Iron(III)

	ligand		
rate constant[a]	Tiron	acetohydroxamic acid[b,c] (ΔV^*, cm^{-3} mol^{-1})[b]	N-methylacetohydroxamic acid[d]
k_1' (M^{-1} s^{-1})	1.7×10^3	5.7×10^3 [c] (5.2)	2.65×10^3
k_1 (M^{-1} s^{-1})	2.3	4.8[c] (−6.3)	1.8
k_2 (M^{-1} s^{-1})	3.2×10^3	2.4×10^3 (−2.2)	0.81×10^3
k_3 (M^{-1} s^{-1})	7.7×10^3	1.6×10^3 (−2.2)	0.75×10^3
k_{-1}' (s^{-1})	1.3	7.8×10^{-2} (−3.3)	7.1×10^{-3}
k_{-1} (M^{-1} s^{-1})	1.2	7.2×10^{-2} [c] (−15.6)	3.2×10^{-3}
k_{-2} (M^{-1} s^{-1})	$>2.8 \times 10^5$	1.3×10^3 (−9.1)	1.0×10^2
k_{-3} (M^{-1} s^{-1})	1.7×10^7	9.3×10^4 (−5.4)	8.6×10^3

[a] Values at 25°C in 1.0 M $NaClO_4$ for Tiron and 2.0 M $NaClO_4$ for the hydroxamates. [b] Values from ref 4. [c] In some cases the rate constants from ref 4 are more than 25% different from those quoted in ref 3; in the latter, these values are $k_1' = 2 \times 10^3$, $k_1 = 1.2$, $k_{-1} = 11 \times 10^{-2}$. [d] Values from ref 3.

Scheme 2



and the hydroxamic acids. The fact that $k_2 \gg k_1$ indicates that one catecholate dianionic ligand or one acetohydroxamate monoanion has similar labilizing effects on substitution on aqueous iron(III). Since k_2 is of the same magnitude as k_1', it appears that these anions have a labilizing effect similar to that of a hydroxide ligand. The k_3 values show that two catecholate ligands increase the reactivity by 5−10 times compared to two acetohydroxamate ligands. The dissociation rate constants (k_{-n}) are about 10^2 larger for Tiron than those of the acetohydroxamates, but in all three systems $k_{-2} \approx 10^5 k_{-1}$ and $k_{-3} \approx 10^2 k_{-2}$.

Brink and Crumblis[15] noted isokinetic relationships for the individual pathways for the reactions of a series of substituted hydroxamic acids with aqueous iron(III). The Tiron values of ΔH^* and ΔS^* for k_1 and k_1' fit reasonably on the appropriate isokinetic relationships of Brink and Crumblis. The latter suggested that these relationships indicate a common mechanism for a particular k_i with some associative character because of the ligand dependence of ΔH^* and ΔS^*. It also is possible that the isokinetic behavior reflects a trend in mechanism, such as increasing associative character with decreasing ΔS^*.

The values of K_{cn} are difficult to categorize because they are potentially a combination of proton release and chelate ring closing. The values decrease with increasing n as expected on simple charge arguments for either proton release or ring closure. However, there is a much larger difference between K_{c2} and K_{c3} than between K_{c1} and K_{c2}. The system can be represented by Scheme 2, where M−LH and M=LH are the protonated monodentate and bidentate species, respectively, and M−L and M=L are the corresponding deprotonated species.

The equilibrium constants in Scheme 2 are related by $K_{rcH}K_{ba} = K_{ma}K_{rc}$, and the experimental K_{cn} is given by eq 7. If the

$$K_{cn} = \frac{([M{-}L] + [M{=}L])[H^+]}{([M{-}LH] + [M{=}LH])} = \frac{K_{ma}(1 + K_{rc})}{1 + K_{rcH}} \qquad (7)$$

ring-closing equilibrium constants K_{rc} and K_{rcH} are much less than 1, then $K_{cn} = K_{ma}$, but if they are much larger than 1, then $K_{cn} = K_{ma}K_{rc}/K_{rcH} = K_{ba}$. If ring closure is very favorable for the deprotonated form, but unfavorable for the protonated form ($K_{rc} \gg 1$, $K_{rcH} \ll 1$), then $K_{cn} = K_{ma}K_{rc}$. As a model for K_{ma}, one might use the observation of Liang and Gould[16] that the −OH group of the salicylate complex of pentamminecobalt(III) has a $pK_a = 9.77$. Ionization of the second−OH in Tiron has a pK_a ~1.3 units larger than that of the −OH in salicylate, and $Fe(OH_2)_6^{3+}$ is ~3 pK units more acidic than $(H_3N)_5Co(OH_2)^{3+}$, so that one might estimate that $pK_{ma} \approx 6$. Therefore it seems unlikely that $K_{c1} = 0.96$ M should be assigned to K_{ma}. Analogies for K_{ba} can be found from the result of Liang and Gould[16] that the chelated salicylate complex $(H_2O)_4Cr(O_2CC_6H_4OH)^{2+}$ has a $K_a \approx 0.4$ and a similar value of ~0.8 M[17] for the salicylate complex $(H_2O)_4Fe(O_2CC_6H_4OH)^{2+}$, determined from the formation and dissociation kinetics. Then $K_{c1} = 0.96$ M obtained in this study seems consistent with assignment to K_{ba}. For K_{c2}, the lower limit of >0.015 M also would seem more consistent with assignment to K_{ba}.

The large difference between K_{c1} and K_{c3} indicates that they probably are for different processes. In the tris complex, the two spectator catecholate ligands should make ring-closing processes less favorable than in the mono complex. Then it seems possible that the conditions leading to $K_{c3} = K_{ma}K_{rc}$ or $K_{c3} = K_{ma}$ are applicable.

Experimental Section

Materials. The solutions of iron(III) perchlorate and Tiron (Eastman) were prepared as described elsewhere,[14] as were the perchloric acid and sodium perchlorate.

Stopped-Flow Studies. The absorbance/wavelength/time spectra (Figure 1) were obtained on an Applied Photophysics SQ.1 instrument. The single-wavelength kinetics were studied by mixing equal volumes of solutions of iron(III) plus Tiron and perchloric acid or acetate buffer on a Tritech Dynamics Model IIA system. In this system, the storage and drive syringes and mixing and observation chambers are all submerged in a water bath to give optimum temperature uniformity. The reagent concentrations and monitoring wavelengths for the different stages are given in the Results section. Each rate constant is the average of eight replicate determinations, evaluated by least-squares fitting to a three-parameter first-order rate law ($A_t = A_\infty - (A_\infty - A_0)\exp(-kt)$) on a DOS 486/33 computer system interfaced to the stopped-flow system.

Numerical Analysis. The nonlinear least-squares program uses a Newton−Raphson procedure[18] with the Marquardt algorithm and has been adapted to run in BASIC on a MacIntosh platform. This program allows the use of more than one independent variable so that data at different temperatures and $[H^+]$ can be combined to obtain activation parameters from one overall fit. The errors quoted are 1 standard deviation.

The constancy of the free Tiron concentration $[H_2L]$ in eqs 1 and 2 is not strictly satisfied. For the mono and tris systems, the $[H_2L]$ is ~8 and ~6 times greater than the iron(III) concentration, but this factor is only 2.6 for the bis system. This is not a significant problem for the latter because the $[H_2L]$ term contributes <15% to k_{exp}. As a result,

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(15) Brink, C. P.; Crumblis, A. L. *Inorg. Chem.* **1984**, *23*, 4708.
(16) Liang, A.; Gould, E. S. *J. Am. Chem. Soc.* **1970**, *92*, 6791.
(17) Jordan, R. B. *Inorg. Chem.* **1983**, *22*, 4160.
(18) Share No. SDA 3094, IBM 360-65.

the coefficients for the [H_2L] term are poorly defined for the bis to mono reaction. It was assumed in the analysis of all systems that [H_2L] = initial free [H_2L] + 0.5[iron(III)]; i.e., [H_2L] has the value at 50% dissociation. This assumption was found to be consistent with numerical integration analysis for the tris to bis reaction under conditions with initial [H_2L]:[iron(III)] ratios from 4 to 10 in which the [H_2L] term contributed from 20 to 40% to k_{exp}. It turns out that there is significant self-correction to this factor. When dissociation is essentially complete, the [H_2L] term makes little contribution; when dissociation is not complete, the [H_2L] changes less.

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Supporting Information Available: The temperature, [H^+], and experimental and calculated rate constants for the mono, bis, and tris complex reactions (Tables S1−S3, respectively) and representative initial species distributions for pH 2−7 (Figure S1) (5 pages). Ordering information is given on any current masthead page.

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