high and a pronounced preference for ortho fluorination is observed with substituted aromatics. In Table II, a partial list of screening reactions that have been carried out with $(CF_3SO_2)_2NF$ under nearly identical conditions are summarized. The reactions are readily monitored by ^{19}F NMR, by observing the growth of aryl-F with time, the disappearance of the *N*-fluorosulfonimide resonances, and the growth of the resonances of the parent sulfonimide. We have carried out reactions of the other sulfonimides with benzene, toluene, and other aromatics, and the observed results are very similar, except that $(CF_3SO_2)N-F$ is clearly the most reactive of the compounds in Table I.

In addition to their utility in direct aromatic fluorinations, these reagents are clearly useful in the fluorination of carbon anions. Reaction of $(CF_3SO_2)NF$ with sodium diethyl 1-methylmalonate in $CDCl_3$ at −10 °C gave a 96% isolated yield of diethyl 1-fluoro-1-methylmalonate along with the expected sodium salt of the imide, $NaN(SO_2CF_3)_2$. This is one of the highest yields that we are aware of for this type of reaction.[13] The only limitation we see for this use is the need for a solvent which will not itself react with the *N*-fluorosulfonimide.

The reactivity shown by these compounds is unique in our opinion. We have not demonstrated that these N-fluorosulfonimides can be used on a preparative scale to prepare a given monofluoro aromatic nor have we defined the scope of these reactions. Such work is in progress, and our earlier statement of potentially wide applicability in selective fluorinations continues to be warranted.

Acknowledgment. The financial support of this research by the National Science Foundation (CHE-8517336) is gratefully acknowledged. We also acknowledge the 3M Company for generous gifts of perfluoroalkanesulfonyl fluorides.

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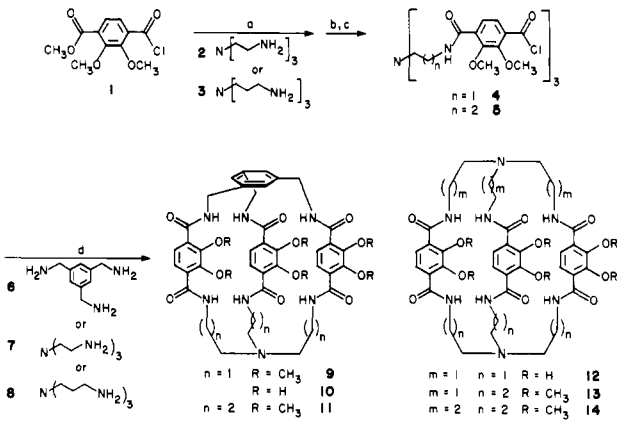

Figure 1. Synthetic scheme: a = amine **2** or **3**, Et_3N, THF, room temperature, 24 h, 80–88%; b = NaOH, CH_3OH, reflux for 2 h, 94–97%; c = $SOCl_2$, THF, DMF, room temperaure, 21 h, intermediate not isolated; d = a solution of amine **6**, **7**, or **8**, Et_3N, THF, and **4** or **5**, high dilution, room temperature, 24 h, 10–15%; e = BBr_3, $CHCl_3$, 0–25 °C, 48 h, 75%.

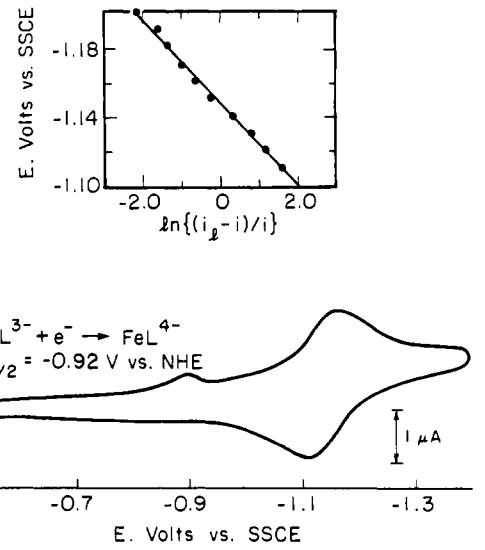

Figure 2. Electrochemistry of Fe(III) **10**. A one compartment cell with Hg as the working electrode, a Pt wire as the counter electrode, and a sodium chloride saturated calomel electrode (SSCE) as the reference electrode were used for all electrochemical studies. The solution was 0.2 mM in the ferric complex of **10**, 0.4 M in $NaClO_4$. The pH 12 was adjusted with KOH. (a) CV performed at a hanging Hg drop electrode and scan rate of 200 mV/s. (b) *E* versus ln $[(i_l - i)/i]$ was plotted by using the curent versus potential trace from NPP performed at 2.0 mV/s, pulse amplitude of 25 mV, and a flow rate of 0.7 mg/s.

Macrobicyclic Iron(III) Sequestering Agents[1]

Thomas J. McMurry, Mir Wais Hosseini,
Thomas M. Garrett, F. Ekkehardt Hahn, Zelideth E. Reyes,
and Kenneth N. Raymond*

Department of Chemistry, University of California
Berkeley, California 94720

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Macrobicyclic ligand designs first appeared in polyether complexing agents (such as the cryptands)[2] and later in polyamines (such as the sepulchrates).[3] These ligand topologies confer remarkable properties on their metal complexes. However, until the recent synthesis of compounds incorporating catechol binding subunits,[4,5] few other macrobicyclic ligands, particularly those designed to complex Fe(III) or other high valent transition metals, had been developed. Such new ligands, while incorporating the three catechol groups found in the siderophore enterobactin,[6] differ significantly in topology from earlier enterobactin analogues[7-12] and are related to phenolic cyclophane[13] macrocycles such as the

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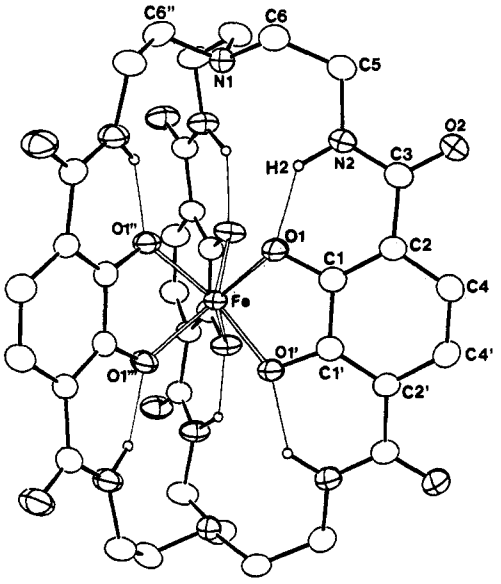

Figure 3. ORTEP of ferric (bicapped TRENCAM). The ellipsoids are scaled to represent the 50% probability surface. Select bond distances (Å) and angles (deg): Fe–O1, 2.012 (2); C1–O1, 1.333 (3); C3–N2, 1.332 (4); N2–H2, 0.89 (4); O1···O1′, 2.526 (4); O1···O1″, 2.713 (3); O1···N2, 2.620 (3); O1···H2, 1.89 (4); O1–Fe–O1′, 77.75 (11); O1–Fe–O1″, 84.78 (8); O1–Fe–O1‴, 134.18 (4); O1···H2–N2, 138 (4).

calixarenes[14] and spherands.[15] Due to the preorganization of their metal binding sites, these ligands are expected to form highly stable complexes that undergo slower ligand exchange kinetics than their acyclic analogues. An examination of the fundamental thermodynamics and kinetics of metal binding requires a set of compounds in which the cavity size, rigidity, and charge is systematically varied. We report here a general series of such compounds, their initial characterization as Fe(III) sequestering agents, and the first structural characterization of any tris(catechoylamide) complex—which also revealed a novel coordination geometry for Fe(III).

The key step in the synthesis (step d, Figure 1) of these ligands is the coupling of two tripodal subunits via a triple amide bond forming reaction. This was performed in THF[16] in the presence of triethylamine under high dilution conditions and afforded the methyl-protected compounds **9**, **11**, **13**, and **14** in 10–15% yield. The tripods **4** and **5** used in this step were made from the known compound **1** in three steps in 78–83% overall yield.

Deprotection of **9** afforded **10** in 75% yield by reaction with BBr_3 in $CHCl_3$. NMR of the Ga(III) complex[17] and positive ion fast atom bombardment (FAB) mass spectrometry confirmed the proposed structure.[18] The compound was crystallized from CH_3OH and gave one peak by HPLC.[19] The ferric complex, obtained by reaction with $FeCl_3$ at pH 10 in H_2O, gives rise to a UV–vis spectrum characteristic of iron complexes of the 2,3-dihydroxyterephthaloyl derivatives,[20] with ligand-to-metal charge-transfer bands at 376, 430, and 528 nm.

Figure 2 shows a typical cyclic voltammogram (CV) of the Fe(III) complex of **10**. The ratio of cathodic-to-anodic peak currents is 1.0, and the peak separation is 60 mV at scan rates of 20–200 mV/s. A plot of E versus ln $[(i_l - i)/i]$ from a normal pulse polarogram (NPP) yields a slope of 24 mV and this together with the CV results indicate a one-electron, mass-transfer-limited process.[21] From a differential pulse polarogram a formal reduction potential of −0.92 V versus NHE was obtained. This corresponds to a ratio of the formation constants for Fe(III)/Fe(II) of $10^{28.6}$.

Crystals of the Fe(III) complex of bicapped TRENCAM (**12**) were obtained from $MeOH/H_2O$. The structure analysis[22] is the first for any ferric tris(catechoylamide) and shows the coordination geometry around the metal center to be trigonal prismatic, a structure previously unknown for Fe(III). Geometries intermediate[23,24] between octahedral and trigonal prismatic have been observed for ferric complexes of bidentate ligands with relatively small bite distances.[25,26] In the present structure, shown in Figure 3, the entire catechoylamide group is planar, and the trans configuration of the amide allows a strong hydrogen bond between the amide proton and the coordinated catechol oxygen. The result is a highly stable ligand structure which incorporates three planar catechoylamide groups and six strong amide hydrogen bonds. To accommodate this ligand geometry, the metal ion lies 0.81 Å out of the catechol plane, giving a 30.79 (12)° dihedral angle between this plane and the plane defined by the iron and the two ligating oxygen atoms. While no structure analysis of ferric enterobactin or an analogue has appeared, a trans geometry for the amide bonds has been predicted by Shanzer and Lifson.[27] As seen in the side view of the complex (Figure 3), both tertiary amines are in the "in" conformation.[28] The molecular and crystallographic site symmetry of the complex is C_{3h}.

The deprotection of compounds **11**, **13**, and **14**, the determination of formal stability constants, and further structural characterization of these complexes are in progress. The incorporation of ions with large octahedral ligand field stabilization energies such as Cr(III) into **12** is expected to force the coordination geometry away from the unusual trigonal prismatic geometry found here for the Fe(III) complex.

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Supplementary Material Available: A description of X-ray structure determination and stereoviews and tables of crystal and data collection details, positional parameters, site occupancy

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(17) ^1H NMR (K_3[Ga(III)(**10**)], 500 MHz, D_2O) δ 2.70 (s, N-CH_2-CH$_2$, 6 H), 3.55 (s, N-CH$_2$-CH_2, 6 H), 4.50 (s, Ar-CH_2-N, 6 H), 6.95 (s, Ar-H, 6 H), 7.61 ns, Ar-H, 3 H).
(18) +FAB-MS: m/e 798 (M + H)$^+$.
(19) HPLC conditions: Rainin Dynamax 8 μm C18 column (4.6 × 250 mm); A = 0.025 M HCO_2H/H_2O, pH 2.69; B = 0.025 M HCO_2H/CH_3OH; ramp 20–100% B over 20 min.
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factors, bond distances, angles, hydrogen bonds and selected nonbonding distances, least-squares planes, anisotropic thermal parameters, and root-mean-square amplitudes of thermal vibration (27 pages); tables of values of F_{obsd} and F_{calcd} (13 pages). Ordering information is given on any current masthead page.

The Ionization of Terminal Acetylenes: Pseudoacid Behavior

T. Aroella, C. H. Arrowsmith, M. Hojatti, A. J. Kresge,* M. F. Powell, Y. S. Tang, and W.-H. Wang

Department of Chemistry
University of Toronto
Toronto, Ontario M5S 1A1, Canada
Received May 26, 1987

The acid ionization of terminal acetylenes, eq 1, generates

$$RC{\equiv}CH \rightleftarrows R{\equiv}C:^- + H^+ \quad (1)$$

carbanions whose basic electron pairs reside in sp orbitals localized on single carbon atoms. Such carbanions thus resemble "normal"[1] oxygen and nitrogen bases, whose basic electron pairs are also localized on single atoms, and acetylenes might therefore be expected to show "normal" acid–base behavior. Recent evidence in support of this idea comes from Brønsted relations for the thermodynamically uphill ionization of phenylacetylene[2] and several other terminal monoalkynes:[3] these Brønsted relations have unit slopes and thus resemble the unit-slope uphill legs of Eigen plots[1] for "normal" acid–base reactions, in which proton transfer is rapid and separation of the proton-transfer products is rate determining, eq 2. We wish to report that we have now

$$B + HA \overset{\text{fast}}{\rightleftharpoons} BH^+{\cdot}A^- \xrightarrow{\text{r.d.}} BH^+ + A^- \quad (2)$$

found other terminal acetylenes whose uphill ionizations give Brønsted relations with slopes decidedly less than unity. This plus additional isotope effect evidence indicates that proton transfer is rate determining in these acid–base reactions; these acetylenes are therefore not functioning as "normal" acids but rather are showing the regular "pseudoacid" behavior typical of most carbon acids.

We measured rates of detritiation of cyanoacetylene in aqueous solution catalyzed by amines and carboxylate ions, eq 3. The

$$CNC{\equiv}CT \xrightarrow[B]{H_2O} CNC{\equiv}CH \quad (3)$$

data give separate Brønsted relations for the two kinds of catalyst, whose slopes are $\beta = 0.65 \pm 0.04$ (amines) and $\beta = 0.83 \pm 0.03$ (carboxylate ions).[4,5] We also examined the detritiation of two conjugated diacetylenes catalyzed by amines, eq 4, and find $\beta = 0.79 \pm 0.04$ (Ar = 4-$CH_3C_6H_4$) and $\beta = 0.79 \pm 0.06$ (Ar = 3-$CH_3OC_6H_4$).[4]

$$ArC{\equiv}C{-}C{\equiv}CT \xrightarrow[B]{H_2O} ArC{\equiv}C{-}C{\equiv}CH \quad (4)$$

We have, in addition, measured rates of dedeuteriation of cyanoacetylene in aqueous formic acid–formate buffers. The results, when combined with detritiation rate constants measured in the same buffer system, give isotope effects which, by way of the relationship $k_H/k_D = (k_D/k_T)^{2.26}$,[7] are equivalent to $(k_H/k_D)_{HCO_2^-} = 5.44 \pm 0.99$ and $(k_H/k_D)_{HO^-} = 2.41 \pm 0.34$. We have also determined $(k_H/k_D)_{HO^-} = 2.13 \pm 0.11$ and $(k_H/k_D)_{HO^-} = 1.48 \pm 0.23$ for hydron transfer from the hydroxide ion to phenylacetylene and chloroform, respectively; chloroform is a carbon acid whose conjugate base also contains a localized electron pair, and whose unit-slope Brønsted relation[2] implies normal acid behavior.

Table I. Analysis of Isotope Effects

substrate	pK_a	base	ϕ^*	$\lvert\Delta pK\rvert^a$
$CHCl_3$	24	HO^-	0.75 ± 0.12	8
PhC≡CH	20	HO^-	0.30 ± 0.02	4
NCC≡CH	8	HO^-	0.27 ± 0.04	8
NCC≡CH	8	HCO_2^-	0.12 ± 0.02	4

[a] Absolute value of pK_a(substrate) - pK_a(base).

It is instructive to analyze these isotope effects in terms of fractionation factor theory.[8] This gives $k_H/k_D = \phi_s/\phi^*$, where ϕ_s is the D:H fractionation factor for the isotopically substituted hydrogen of the substrate in the initial state and ϕ^* is the fractionation factor of this hydrogen in the transition state. The value of ϕ_s for chloroform is 1.11 ± 0.01,[9] and use of that with $k_H/k_D = 1.48$ leads to $\phi^* = 0.75 \pm 0.12$. Values as close to unity as this are not characteristic of hydrogens undergoing transfer, but this value of ϕ^* is reasonable for the hydrogen atom of a water molecule solvating the negative charge of an already fully formed trichloromethyl carbanion, as shown in the product of eq 5; for

$$Cl_3CL + HO^- \rightarrow Cl_3C:^-LOH \quad (5)$$

example, $\phi = 0.7$ for the hydrogens of a water molecule performing the same function in the solvated hydroxide ion.[8c,10] Similar analysis of the isotope effects on ionization of the acetylenes, using $\phi_s = 0.64$ for the acetylenic hydrogen,[8c,11] gives $\phi^* = 0.30$ to 0.12 (Table I). Such low values *are* characteristic of hydrogens undergoing transfer; they represent sizable primary kinetic isotope effects.

This difference between chloroform and the acetylenes might be a consequence of differences in the energetics of these reactions: the proton-transfer step of energetically balanced systems with ΔpK near zero will be slower than that of imbalanced systems with ΔpK substantially removed from zero,[12] and the difference may be sufficient to change normal acid to pseudoacid behavior. Examination of this hypothesis requires information on the pK_a's of the present substrates. Directly measured values for aqueous solution are not available, but the estimates $pK_a = 24$ for chloroform and $pK_a = 20$ for phenylacetylene have been made from hydrogen-exchange rates.[2] Use of this method for cyanoacetylene can give only an upper limit, $pK_a < 13$, but this is enough to place the pK_a of this substance between that of the conjugate acids of the two bases, HO^- and HCO_2^-, used here to determine isotope effects; the assumption that these isotope effects vary in magnitude with ΔpK according to the Melander–Westheimer principle[13] then leads to $pK_a = 8$. These estimates give the ΔpK values listed in Table I.

It may be seen that the normal acid behavior indicated by $\phi^* = 0.75$ for proton transfer from chloroform to hydroxide ion is indeed associated with a greater value of ΔpK than the pseudoacid

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