the χ values for all ortho azanucleosides (8-azapurine or 6-azapyrimidine nucleosides)[4,5,7,10-15] is presented in Table I and shows that with one exception they lie either in the intermediate "high-anti" region ($\chi_{mean} \approx 100$) or in the syn region. The lone exception is that one of the two known forms of virazole[13] is in the anti conformation while the other adopts the "high-anti" conformation; virazole, however, has much more conformational freedom than the other nucleosides listed in Table I. It should be noted, however, that the syn conformation for the two pyrazomycins is stabilized by strong intramolecular hydrogen bonding between a base hydroxyl and O(5′) in pyrazomycin A (the β-anomer) and the same hydroxyl and O(2′) in pyrazomycin B (the α-anomer). In formycin hydrobromide the formycin cation is protonated at N(1). It is clear, however, that the unsubstituted ortho azanucleosides favor the intermediate "high-anti" region rather than the syn region. The main intramolecular interactions in those 8-azapurine nucleosides which exist in the "high-anti" χ region are between N(8) of the base and C(2′) of the sugar (and its attached hydrogen H(2′)) (see Table I); in 8-azaadenosine, the N(8)···C(2′) and N(8)···H(2′) separations are 2.835 (7) and 2.34 (5) Å, respectively, and the N(8)···H(2′)–C(2′) angle is approximately 100°. This is similar to the situation found for 6-azapyrimidine nucleosides, 6-azauridine[10] and 6-azacytidine,[11] where two of the severest intramolecular interactions are between the same two sugar atoms, namely, C(2′) and H(2′), and N(6) of the base.

Sundaralingam[16] has shown that the glycosidic bond in purine nucleosides (average value 1.465 Å) is shorter than that in pyrimidine nucleosides (1.495 Å). A comparison of the glycosidic bond lengths for 8-azaadenosine (1.445 Å) with those of 6-azauridine (1.468 Å)[10] and 6-azacytidine (1.465 Å)[11] suggests that this trend carries over to the ortho azanucleosides but that the bonds are shorter in the latter group than in the normal nucleosides. This general shortening of the glycosidic bond in ortho azanucleosides may not be related to the ortho-aza substitution, since tubercidin does not have this feature, but instead may be a function of the "high-anti" conformation around the glycosidic bond as suggested by Sundaralingam.[17,18]

Other stereochemical parameters of interest in 8-azaadenosine are similar to those in formycin. Thus, the puckering of the ribose ring is C(2′)-endo, C(1′)-exo, $^2T_1$, which, as noted by Sundaralingam, is slightly different from that normally observed ($^2T_3$) in purine nucleosides. The conformation of the C(5′)–O(5′) bond around the extracyclic bond C(4′)–C(5′) is gauche-gauche as in formycin. The hydrogen bonding scheme is also similar to that of formycin with the exception that N(7), which is not protonated in 8-azaadenosine, does not form the donor hydrogen bond N(7)–H···O-(3′) found in formycin. In 8-azaadenosine N(7) does not participate in any hydrogen bonding.

The observed "high-anti" (intermediate between syn and anti) conformation of 8-azaadenosine may explain its weakened binding (relative to adenosine) to adenosine deaminase,[6] since this enzyme is inactive on nucleosides in the syn conformation but active on nucleosides in the anti conformation.[19]

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Retardation of the Ferric Ion Catalyzed Decomposition of Hydrogen Peroxide by an Iron(II) Diimine Complex

Sir:

Much interest has been devoted to the study of the ferric ion catalyzed decomposition of hydrogen peroxide,[1,2] and also to the retardation of this reaction by organic substrates.[3]

We now report a striking retardation of the ferric ion catalyzed decomposition of hydrogen peroxide by an iron(II) complex,[4] tris(glyoxal bis(methylimine))iron-(II), $Fe(GMI)_3^{2+}$. While in the retardation of the reaction by organic substrates evidence was presented that the organic compound acts as an HO· radical trap,[3] in this case we believe that $Fe(GMI)_3^{2+}$ acts as an HO_2· radical trap. In the reaction of hydrogen peroxide with other diimine complexes of iron(II), such as tris(1,10-phenanthroline)iron(II) and tris(2,2′-bipyridine)iron(II),[2,5] the complexes undergo dissociation and the respective bis and mono complexes of iron(III) catalyze the decomposition of hydrogen peroxide,[6,7] contrary to what is observed in the present case.

It is known that $Fe(GMI)_3^{2+}$ can only be oxidized reversibly to $Fe(GMI)_3^{3+}$ at high acid concentration[8] (*e.g.*, 10 *M* H_2SO_4). At lower acid concentration, disproportionation reactions of the ferric complex occur, leading to the formation of ligand-oxidized complexes.[9] The formal electrode potential of the couple $Fe(GMI)_3^{3+}|Fe(GMI)_3^{2+}$ is 1.02 V *vs.* sce in 4.0 *M* H_2SO_4 and increases as the acid concentration decreases. This behavior is similarly observed in the

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aromatic diimine complexes.[10,11] The compound Fe(GMI)$_3^{2+}$ can be oxidized by strong oxidizing agents (*e.g.*, Ce(IV)).

We now present some evidence for the reaction between $HO_2\cdot$ radicals and Fe(GMI)$_3^{2+}$. In Figure 1 the changes in concentration of hydrogen peroxide (Figure 1a), complex (Figure 1b), and evolution of oxygen (Figure 1c) with time are shown. The rate of decomposition of hydrogen peroxide decreases as the concentration of complex increases, and Fe(III) decreases. When all the colored species are destroyed, the decomposition of hydrogen peroxide catalyzed by iron(III) is observed. The rate of oxygen evolution decreases as the concentration of complex increases. This slower rate of oxygen evolution can be easily explained in the light of the mechanism proposed by Barb, *et al.*,[12] for the decomposition of hydrogen peroxide by Fe(III) ions.

$$\text{Fe(III)} + H_2O_2 \longrightarrow \text{Fe(II)} + HO_2\cdot + H^+ \quad (1)$$

$$\text{Fe(II)} + H_2O_2 \longrightarrow \text{Fe(III)} + HO\cdot + HO^+ \quad (2)$$

$$HO\cdot + H_2O_2 \longrightarrow HO_2\cdot + H_2O \quad (3)$$

$$HO_2\cdot + \text{Fe(III)} \longrightarrow H^+ + \text{Fe(II)} + O_2 \quad (4)$$

$$HO_2\cdot + \text{Fe(II)} \longrightarrow HO_2^- + \text{Fe(III)} \quad (5)$$

By this mechanism, the reaction between $HO_2\cdot$ radicals and Fe(III) is responsible for the oxygen evolution (eq 4). To slow the rate of oxygen evolution, the complex must compete with Fe(III) for $HO_2\cdot$ radicals. Thus the complex is oxidized to Fe(GMI)$_3^{3+}$, regenerating hydrogen peroxide (*cf.* eq 5) and acting as an $HO_2\cdot$ radical trap. The ferric complex so formed can undergo an internal redox reaction[9] producing a radical at the ligand. The complex–radicals[13] which are generated may react with $HO_2\cdot$ or $HO\cdot$ radicals to form the labile ligand-oxidized species.

As the formal electrode potential of the couple $HO_2\cdot|H_2O_2$ is approximately 1.5 V,[14] it is reasonable to assume that the oxidation of Fe(GMI)$_3^{2+}$ is by $HO_2\cdot$ radicals. This reaction must be very fast since the rate of electron exchange in analogous systems, *e.g.*, between Fe(phen)$_3^{2+}$ and Fe(phen)$_3^{3+}$ is very high[15] ($3 \times 10^8\ M^{-1}\,\text{sec}^{-1}$) (*cf.* Figure 1b).

Attempts to oxidize Fe(GMI)$_3^{2+}$ with $HO\cdot$ radicals gave a very low yield of the labile ligand-oxidized complexes. However, these experiments are not completely analogous since the acid concentration is much higher (0.5 *M* H_2SO_4) than that of the iron(III) experiments (pH 2.4). These radicals were generated by treating hydrogen peroxide with excess iron(II)[16] (in the presence of complex).

The rate laws deduced, including a reaction between the complex and $HO\cdot$ radicals and the reactions of the Barb, *et al.*, mechanism,[12] seem not to be compatible with our preliminary kinetic data. This observation indicates that a simple reaction between the complex and $HO\cdot$ radicals is not likely to be in the main reaction path.

The kinetics of this reaction are being carefully measured. The effect of Fe(GMI)$_3^{2+}$ and of other diimine complexes on other catalysts of the hydrogen peroxide decomposition is under investigation.

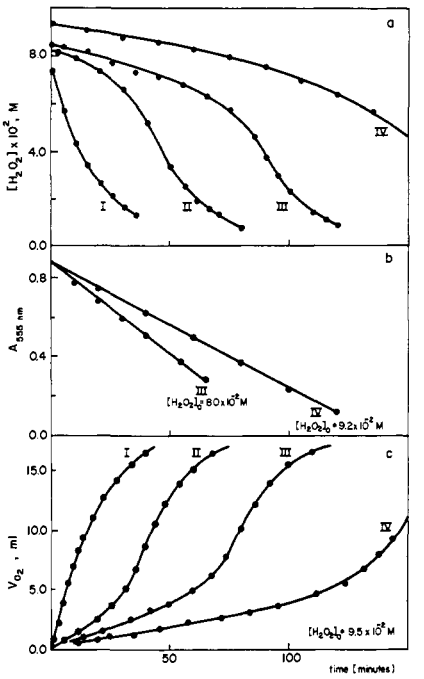

Figure 1. Plots of (a) hydrogen peroxide concentration *vs.* time, (b) absorbance at 555 nm (absorption maximum of Fe(GMI)$_3^{2+}$, $\epsilon_{555} = 8.8 \times 10^3$) *vs.* time, and (c) volume of oxygen evolved *vs.* time. All data refer to reactions at 35.0° and pH 2.4 (H_2SO_4). Concentration of iron(III) is the same in experiments I, II, and III ($4.2 \times 10^{-3}\ M$), and $2.1 \times 10^{-3}\ M$ in experiments IV. Concentration of Fe(GMI)$_3^{2+}$ is the same in experiments III and IV ($1.0 \times 10^{-4}\ M$) and $5.0 \times 10^{-5}\ M$ in experiment II. Experiments I contain no Fe(GMI)$_3^{2+}$.

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Nuclear Magnetic Resonance Studies of Hemoproteins. Unusual Temperature Dependence of Hyperfine Shifts and Spin Equilibrium in Ferric Myoglobin and Hemoglobin Derivatives[1]

Sir:

Nuclear magnetic resonance (nmr) spectroscopy of paramagnetic macromolecules has recently developed into a powerful tool for investigating structure and structure–function relationships in metalloproteins such

(1) A part of our systematic investigation on nuclear magnetic resonance studies of hemoproteins and hemoenzymes.