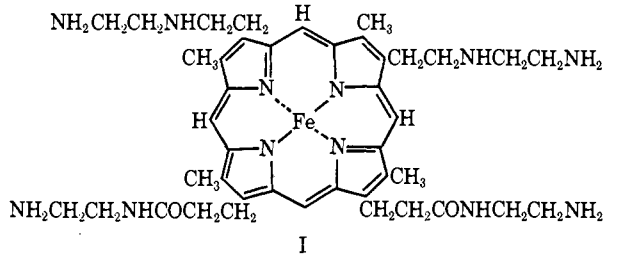
Temperature-Jump Investigation of the Kinetics of Imidazole Substitution on an Iron(III) Porphyrin in Aqueous Solution[1a]

Gerald B. Kolski*[1b] and Robert A. Plane

Contribution from the Department of Chemistry, Cornell University, Ithaca, New York 14850. *Received August 27, 1971*

Abstract: An investigation of the reactions of the iron(III) derivative of the ethylenediamine-substituted protoporphyrin IX in aqueous solution gave hydrolysis constants for the iron(III) porphyrin of $K_1 = 3 \times 10^{-5}$ M and $K_2 = 1 \times 10^{-6}$ M for the reactions $FeP(H_2O)_2^+ \rightleftarrows FeP(H_2O)(OH) + H^+$ and $FeP(H_2O)(OH) \rightleftarrows FeP(OH)_2^- + H^+$. The dimerization constant for the reaction $2FeP(OH)_2^- \rightleftarrows O(FeP(OH))_2^{2-}$ was found to be 1×10^5 M^{-1} at 25° and 0.1 M KNO_3. The rate of substitution of water by imidazole for the aquo iron(III) porphyrin was found to be $k_{12} = (0.85 \pm 0.15) \times 10^7$ M^{-1} sec^{-1}. The reverse rate constant for the dissociation of the imidazole complex was $k_{21} = (2.2 \pm 0.4) \times 10^3$ sec^{-1}. The equilibrium constant determined kinetically was in good agreement with a value of $K_{12} = (3.8 \pm 1.3) \times 10^3$ M^{-1} determined spectrophotometrically. The results are discussed in terms of a mechanism and compared with the known chemical behavior of Fe(III). The presence of the porphyrin decreases the acidity of Fe(III) a thousandfold and increases the lability a thousandfold.

Previous studies of ferriporphyrins in aqueous solutions have often been hampered by the low solubility of the iron(III) porphyrins at low pH.[2,3] The iron(III) porphyrin shown in I was soluble throughout the pH range from 1 to 10. This enabled a thorough investigation of the hydrolytic behavior of the iron(III) porphyrin and its effect on the substitution of imidazole in the out-of-plane positions. Unlike previous studies[4,5] the identity of the reacting hydrolytic species



I

could be established. The reactions shown in eq 1 and 2 were characterized and found to be fast enough to

$$FeP(H_2O)_2^+ + Imid \rightleftarrows FeP(Imid)(H_2O)^+ + H_2O \quad (1)$$

$$FeP(OH)(H_2O) + Imid \rightleftarrows FeP(Imid)(OH) + H_2O \quad (2)$$

observe using temperature-jump relaxation methods.

Experimental Section

The ethylenediamine-substituted protoporphyrin IX was prepared and purified as previously described.[6,7] The iron(III) porphyrin was prepared by the addition of an excess of ferrous chloride to a refluxing solution of the porphyrin in *N,N'*-dimethylformamide (DMF).[8] The reaction was considered complete when a drop of the reaction mixture added to 10^{-2} M HCl gave no spectral indication of a band at 403 nm. This band is characteristic of the porphyrin species having all four pyrrole nitrogens protonated and is readily observable in the presence of ferriporphyrin. The DMF was removed by evaporation. The solid which remained was dissolved in H_2O and the pH adjusted to 8 with sodium hydroxide to remove the excess iron as ferric hydroxide. The ferriporphyrin was then salted out with NaClO and dissolved in H_2O, and the ClO_4^- exchanged for Cl^- on an anion exchange column of Bio Rex 9, 100–200 mesh.

Temperature-jump experiments were performed on the instrument previously described.[6] The measurements were made at ~1.1 M ionic strength maintained with added potassium nitrate. The reaction was observed in the Soret region at 405 nm near the maximum for the ferriporphyrin–imidazole complex. The imidazole was standardized by potentiometric titration.

Spectrophotometric measurements were made on a Cary 14 or a Bausch and Lomb 505 spectrophotometer. The spectrophotometric determination of the hydrolysis constants of the ferriporphyrin complex and the stability constant of the ferriporphyrin–imidazole complex was carried out at 25° using a thermostated cell for the Cary 14. The stability constant for the imidazole–ferriporphyrin complex was determined using the spectrophotometric data and the equation

$$K_{12} = \frac{(A/[FeP_T])(1 + K_1/[H^+]) - (\epsilon_1 + \epsilon_2 K_1/[H^+])}{(\epsilon_3 - (A/[FeP_T]))[Imid]}$$

where A is the absorbance, $[FeP_T]$ is the total ferriporphyrin concentration, K_1 is the first hydrolysis constant for the iron(III) porphyrin, [Imid] is the concentration of unprotonated imidazole, and ϵ_1, ϵ_2, and ϵ_3 are the molar absorptivities for the $FeP(H_2O)_2^+$, $FeP(OH)(H_2O)$, and $FeP(H_2O)(Imid)^+$ species, respectively, at 390 nm. The values of the molar absorptivities are $\epsilon_1 = 1.03 \times 10^5$ M^{-1} cm^{-1}, $\epsilon_2 = 6.5 \times 10^4$ M^{-1} cm^{-1}, and $\epsilon_3 = 4.4 \times 10^4$ M^{-1} cm^{-1}.

The values of the molar absorptivities ϵ_1 and ϵ_2 and the hydrolysis constant K_1 were obtained from an independent study of the hydrolytic behavior of the ferriporphyrin. The molar absorptivity ϵ_3 of the imidazole complex was obtained from studies where the complex was fully formed.

The hydrolysis constants K_1 and K_2 were evaluated in regions where the parameter $A/[FeP_T]$ was independent of porphyrin concentration.

$$A/[FeP_T] = \frac{(\epsilon_1 + \epsilon_2 K_1/[H^+] + \epsilon_4 \beta_2/[H^+]^2)}{(1 + K_1/[H^+] + \beta_2/[H^+]^2)}$$

The molar absorptivity for $FeP(OH)_2^-$ is $\epsilon_4 = 3 \times 10^4$ M^{-1} cm^{-1}. These constants were used to evaluate the dimerization constant, K_D, in the pH range where there was a definite porphyrin concentration dependence.

Beers Law. Since previous results using the ethylenediamine-substituted protoporphyrin IX have shown evidence for dimerization of the free porphyrin,[6] steps were taken to determine if this occurred with the iron(III) derivative of this porphyrin at low pH. A Beer's law plot of the Soret at 390 nm for the iron(III) porphyrin

(1) (a) Supported by a research grant from the National Institutes of Health, U. S. Public Health Service. (b) NIH Postdoctoral Fellow.
(2) W. A. Gallagher and W. B. Elliott, *Biochem. J.*, **108**, 131 (1968).
(3) D. Keilin, *Proc. Roy. Soc., Ser. B*, **100**, 129 (1926).
(4) B. B. Hasinoff, H. B. Dunford, and D. G. Horne, *Can. J. Chem.*, **47**, 3225 (1969).
(5) E. B. Fleischer, S. Jacobs, and L. Mestichelli, *J. Amer. Chem. Soc.*, **90**, 2527 (1968).
(6) R. R. Das, R. F. Pasternack, and R. A. Plane, *ibid.*, **92**, 3312 (1970).
(7) T. P. Stein and R. A. Plane, *ibid.*, **91**, 607 (1969).
(8) A. Adler, F. R. Longo, F. Kampas, and J. Kim, *J. Inorg. Nucl. Chem.*, **32**, 2443 (1970).

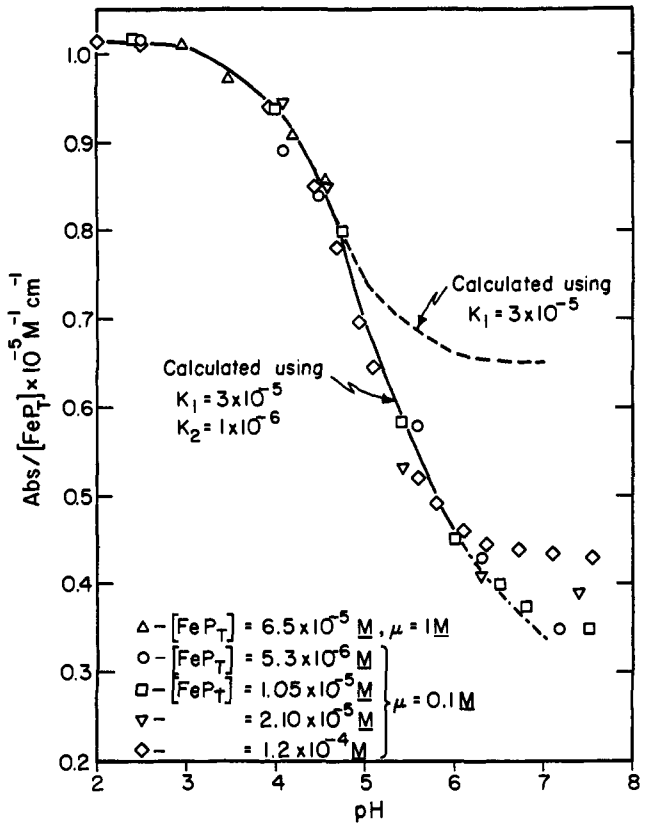


Figure 1. Determination of hydrolysis constants.

showed linearity from 10^{-7} to 10^{-4} M at pH 2. The Soret is extremely sensitive to aggregation unlike some of the other visible peaks.[9] The molar absorptivity ϵ_1 at 390 nm obtained from this plot was 1.03×10^5 M^{-1} cm^{-1}. Aggregation was also checked by adding ethanol to the iron(III) porphyrin solutions since ethanol is known to inhibit aggregation of porphyrins.[6] Since the absorption spectrum was unaffected by this addition, dimerization at this pH was ruled out.

Results

Whereas the iron(III) porphyrin solutions obeyed Beer's law reasonably well below pH 2, at pH above 5, there was considerable deviation from linearity. The curve in Figure 1 shows the effect of pH on the absorbance at 390 nm. The reaction occurring below pH 5 is independent of the concentration of ferriporphyrin and is fully explainable by reaction 3. The hydrolysis

$$FeP(H_2O)_2^+ \overset{K_1}{\rightleftharpoons} FeP(H_2O)(OH) + H^+ \quad (3)$$

constant K_1 has a value of 3×10^{-5} M. As the pH is increased an equilibrium is established which is dependent on the concentration of ferriporphyrin. The aggregation is not that of $FeP(H_2O)(OH)$ but of a more hydrolyzed complex. The data in Figure 1 can be explained if reactions 4 and 5 are postulated. The constants obtained are $K_2 = 1 \times 10^{-6}$ M and $K_D \approx 1 \times 10^5$ M^{-1}; these were evaluated using an iterative process on a computer and the equations previously described.

$$(H_2O)FeP(OH) \overset{K_2}{\rightleftharpoons} FeP(OH)_2^- + H^+ \quad (4)$$

$$2FeP(OH)_2^- \overset{K_D}{\rightleftharpoons} O(FeP(OH))_2^{2-} + H_2O \quad (5)$$

(9) G. B. Kolski and R. A. Plane, submitted for publication in *J. Amer. Chem. Soc.*

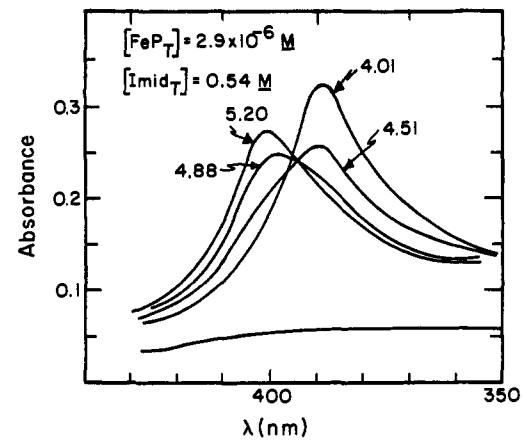


Figure 2. Spectral change for iron(III) porphyrin–imidazole association with pH showing increase in complexation with increase in free imidazole.

Reaction with Imidazole. The reaction of imidazole with the ferriporphyrin was studied at pH below 5 where $FeP(OH)_2^-$ and $O(FeP(OH))_2^{2-}$ were not present in appreciable concentrations in order to simplify the study of the kinetics of formation of the imidazole complex. As Figure 2 shows, there is an appreciable shift in the Soret on addition of imidazole. This shift in Soret was used to study the formation of the imidazole–ferriporphyrin complex. Assuming the reactions shown in eq 6–8, a stability constant for the ferriporphyrin imidazole complex can be calculated. Table I shows the experimental K_{12} calculated from the spec-

$$FeP(H_2O)_2^+ + Imid \overset{K_{12}}{\rightleftharpoons} FeP(H_2O)(Imid)^+ + H_2O \quad (6)$$

$$(H_2O)FeP(OH) + Imid \overset{K_{34}}{\rightleftharpoons} FeP(Imid)(H_2O)^+ + OH^- \quad (7)$$

$$FeP(H_2O)_2^+ \overset{K_1}{\rightleftharpoons} FeP(OH)(H_2O) + H^+ \quad (8)$$

Table I. Determination of Imidazole Association Constant[a]

[FeP$_T$] × 10^5 M	[Imid$_T$], M	pH	$K_{12} \times 10^{-3}$ M^{-1}
3.3	1.14	3.12	3.7
3.3	1.14	3.40	3.5
6.5	1.14	3.50	4.5
3.5	1.14	3.50	1.5
6.5	1.14	3.65	5.2
2.1	1.14	3.70	4.2
3.4	0.76	3.81	3.9
3.5	1.14	3.85	2.5
6.5	1.14	3.90	3.5
2.1	1.14	3.93	3.8
3.3	1.14	3.96	6.2
3.4	0.76	4.10	3.8
3.5	1.14	4.10	3.0
6.5	1.14	4.12	4.4
3.0	0.38	4.26	2.0
2.8	0.23	4.27	1.4
6.5	1.14	4.38	5.0
3.4	0.76	4.48	5.0
2.8	0.23	4.51	2.4
3.4	0.76	4.62	4.2
3.3	1.14	4.82	4.5
2.8	0.23	4.88	4.1
3.0	0.38	4.92	4.6
3.4	0.76	5.10	2.0
2.8	0.23	5.20	5.8
			Av (3.8 ± 1.3)

[a] μ = 1.1 M, temperature = 25°.

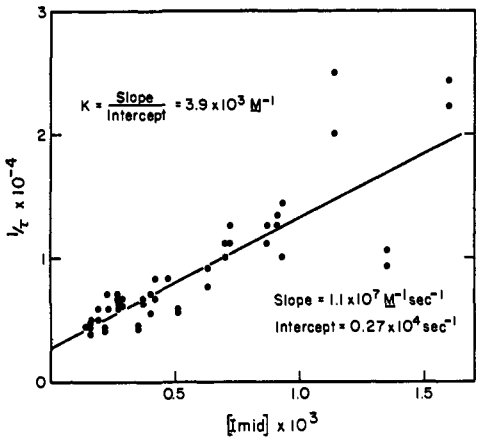


Figure 3. Rate change with free imidazole concentration.

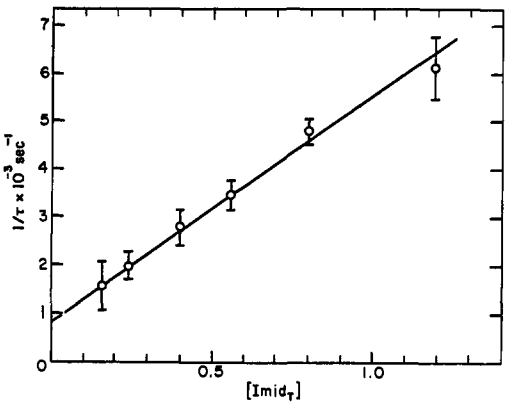


Figure 4. Change in $1/\tau$ with total imidazole at pH 4.0. $[FeP_T] = 6.5 \times 10^{-5}$ *M*.

trophotometric data, $(3.8 \pm 1.3)10^3\ M^{-1}$. Attempts to calculate a constant for the formation of only a 2:1 imidazole–ferriporphyrin complex using the same data gave erratic results with values ranging from $2.4 \times 10^6\ M^{-2}$ to $32.1 \times 10^6\ M^{-2}$. It is evident that a 1:1 complex is more consistent with the data.

The large shift in Soret on complexation also affords a ready means to study the rate of formation of the complex using temperature-jump relaxation methods. Table II indicates the range of relaxation times obtained with changes in total ferriporphyrin and total imidazole concentrations. In the pH range 3.5–5.0 imidazolium ion is the predominant species. There does not, however, appear to be an effect attributable to total imidazole as evidenced by the data. Figure 3, however, shows the dependence of the relaxation time on free imidazole. This would seem to indicate that free imidazole is the predominant reactant. A plot of $1/\tau$ against total imidazole at constant pH is linear as shown in Figure 4. This is consistent with free imidazole being the reactant and also with the formation of only a 1:1 imidazole–ferriporphyrin complex. Using the accepted methods for determining the concentration dependence of the relaxation times[10,11] and reactions 9 and 10 the relationship between $1/\tau$ and the concentration of imidazole and ferriporphyrin was derived for Schemes I and II.

Table II. Data for Imidazole Substitution[a]

$[FeP_T] \times 10^5\ M$	$[Imid_T]$, *M*	$[Imid] \times 10^4\ M$	pH	τ, μ sec
2.8	0.23	1.6	4.15	205
2.8	0.23	2.2	4.27	235
2.8	0.23	3.7	4.51	155
2.8	0.23	5.1	4.65	175
2.8	0.23	8.7	4.88	85
2.8	0.23	11.4	5.00	45
3.0	0.38	1.6	3.92	200
3.0	0.38	3.5	4.26	225
3.0	0.38	9.3	4.70	85
3.0	0.38	16.0	4.92	43
3.4	0.76	1.6	3.63	240
3.4	0.76	2.9	3.88	157
3.4	0.76	6.3	4.22	120
3.4	0.76	13.5	4.55	107
2.1	1.14	2.7	3.68	155
2.1	1.14	4.7	3.92	120
3.3	1.14	2.7	3.82	155
3.3	1.14	9.1	4.20	77
3.4	1.14	1.9	3.52	185
3.4	1.14	4.0	3.85	160
3.4	1.14	7.2	4.10	85
6.5	1.14	1.4	3.50	220
6.5	1.14	2.3	3.64	155
6.5	1.14	4.2	3.90	135
6.5	1.14	7.0	4.12	95

[a] $\mu = 1.1\ M$, temperature = 25°.

$$ImidH^+ \overset{K_a}{\rightleftharpoons} Imid + H^+ \quad \text{fast} \tag{9}$$

$$FeP(H_2O) \overset{K_1}{\rightleftharpoons} FeP(OH) + H^+ \quad \text{fast} \tag{10}$$

Scheme I

$$Imid + FeP(H_2O)_2^+ \underset{k_{21}}{\overset{k_{12}}{\rightleftharpoons}} FeP(Imid)(H_2O)^+ + H_2O$$

$$Imid + FeP(OH)(H_2O) \underset{k_{43}}{\overset{k_{34}}{\rightleftharpoons}} FeP(Imid)(H_2O)^+ + OH^-$$

Scheme II

$$Imid + FeP(H_2O)_2^+ \underset{k'_{21}}{\overset{k'_{12}}{\rightleftharpoons}} FeP(Imid)(H_2O)^+ + H_2O$$

$$Imid + FeP(OH)(H_2O) \underset{k'_{43}}{\overset{k'_{34}}{\rightleftharpoons}} FeP(Imid)(OH) + H_2O$$

The only major difference in these two mechanisms is the reaction of imidazole with the hydroxy complex. One of the mechanisms involves displacement of H_2O and the other displacement of OH^- from the ferriporphyrin–hydroxy complex. The equations for $1/\tau$ are eq 11 for Scheme I

$$1/\tau = k_{12}\{[Imid]\beta + [FeP^+]/(1 + \alpha)\} + k_{34}\{[Imid](1 - \beta) + [FeP(OH)]/(1 + \alpha) + [OH^-]/K_{34}\} + k_{21} \tag{11}$$

where

$$\alpha = \frac{\{[H^+](K_1 + [FeP(OH)] + [H^+]) - [H^+][Imid]\}}{(2[H^+][Imid] + K_1[Imid])}$$

$$\beta = ([H^+] - \alpha[Imid])/[Imid](1 + \alpha)$$

(10) M. Eigen and L. DeMaeyer, "Technique of Organic Chemistry," Vol. VIII, Part II, A. Weissberger, Ed., Interscience, New York, N. Y., 1963, pp 895–1054.
(11) G. W. Castellan, *Ber. Bunsenges. Phys. Chem.*, 67, 898 (1963).

and eq 12 for Scheme II

$$1/\tau = k'_{12}\left\{\frac{[FeP^+]}{(1+a)} + \frac{b[Imid]}{(1+a)} + \frac{b}{K_{12}(1+a)}\right\} + k'_{34}\left\{\frac{[FeP(OH)]}{(1+a)} + \frac{[Imid](1-c-a)}{(1+a)} + \frac{(1+c-b)}{K_{34}(1+a)}\right\} \quad (12)$$

where

$$a = ([H^+]/[Imid])\{[Imid](K_H - K_1) - [H^+](K_1 + [H^+] + K_H + [FeP(OH)] + [FeP(Imid)(OH)]) - K_1K_H - [FePOH]K_H - K_1[FeP(OH)]\}/(-2K_H[H^+] - K_1K_H - [H^+]^2)$$

$$b = \left\{\frac{[Imid](1+a) + [FeP(Imid)(OH)]}{(K_2 + [H^+])}\right\}\frac{[H^+]}{[Imid]}$$

and

$$c = ([H^+]/[Imid]) + b - a$$

The constant K_H, the hydrolysis constant for the imidazole complex

$$K_H = \{[FeP(Imid)(OH)][H^+]/[FeP(Imid)]\}$$

was estimated to be $K_H \leq 3 \times 10^{-6}$ M. This estimate was made from the observation that the spectrophotometrically determined stability constant for the imidazole complex did not vary considerably with pH. Determination of the precise value was hampered by the hydrolysis equilibria for the ferriporphyrin. The concentration dependence for the forward reactions for both schemes is approximately the same in that the imidazole terms dominate and account for the dependence in Figure 3. Rate constants for both schemes are presented in Table III.

Table III. Rate Constants[a]

Scheme I
$k_{12} = (1.0 \pm 0.2) \times 10^7$ M^{-1} sec^{-1}
$k_{34} = (6.0 \pm 0.8) \times 10^6$ M^{-1} sec^{-1}
$k_{21} = (2.5 \pm 0.4) \times 10^3$ sec^{-1}
$k_{43} = (4.6 \pm 1.0) \times 10^{12}$ M^{-1} sec^{-1}
Scheme II
$k'_{12} = (0.85 \pm 0.15) \times 10^7$ M^{-1} sec^{-1}
$k'_{34} = (0.89 \pm 0.08) \times 10^7$ M^{-1} sec^{-1}
$k'_{21} = (2.2 \pm 0.4) \times 10^3$ sec^{-1}
$k'_{43} \approx 2 \times 10^4$ sec^{-1}

[a] $\mu = 1.1$ M, temperature = 25°.

Discussion

Schemes I and II for the imidazole substitution reaction cannot be distinguished using the kinetic data available. The lack of pH dependence for the spectrophotometrically determined equilibrium constant would be consistent with Scheme I. It would also be consistent within experimental error with Scheme II if K_{34}/K_{12} was ≤ 0.1. A comparison of the rate constants evaluated for the two mechanisms shows that the formation rate constants k_{12}, k_{34} and k'_{12}, k'_{34} do not change appreciably with the change in mechanism. The same can be said for the reverse rate constants k_{21} and k'_{21}. The major change occurs in the reverse rate constants k_{43} and k'_{43} on going from the second-order reaction of Scheme I to the first-order reaction of Scheme II. Scheme I would require $k_{43} = 4.6 \times 10^{12}$ M^{-1} sec^{-1}, a constant which is too large for a reaction involving the breaking of the imidazole–iron bond. The rate constant $k'_{43} \approx 2.3 \times 10^4$ sec^{-1} is more consistent with the expected rate of imidazole–iron bond breakage. The value for k'_{43} being approximately an order of magnitude greater than the value for k'_{21} is also consistent with an expected rate enhancement for loss of imidazole from the ferriporphyrin containing a hydroxide bound trans to the imidazole.

The fact that imidazole is more reactive than imidazolium ion in reactions with ferriporphyrin is best explained by a mechanism in which an outer sphere association is followed by the loss of a water from the inner coordination sphere of the ferriporphyrin.

$$FeP(H_2O)_2^+ + Imid \overset{K_{os}}{\rightleftharpoons} FeP(H_2O)_2^+ \cdot Imid \overset{k_{-H_2O}}{\rightleftharpoons} FeP(H_2O)(Imid)^+ + H_2O$$

The charge on the iron(III) porphyrin will be +7 arising from protonation of six nitrogens on the ethylenediamine side chains and the iron(III) and P^{2-}. This would severely limit any outer sphere association between imidazolium and the iron(III) porphyrin and favor association between neutral imidazole and the iron(III) porphyrin. This mechanism would require that the value of k_{-H_2O} be equal to k'_{12}/K_{os} or 9×10^6 M^{-1} sec$^{-1}/K_{os}$. An estimate can be made for K_{os} from the modified Fuoss equation for ion-pair formation[12]

$$K_{os} = (4/3)\pi a^3 N_A e^b \times 10^{-3} \quad (13)$$

where

$$b = |Z_A Z_B| e_0^2/a\epsilon kT$$

In these equations, a is the center-to-center distance (in centimeters) between the two reaction partners at the point of closest approach, N_A is Avogadro's number, Z_A and Z_B are the formal charges on the reacting species, e_0 is the electronic charge (in esu), ϵ is the dielectric constant, k is the Boltzmann constant (in ergs), and T is the absolute temperature.

For the association of a neutral species with a charged species, b can be approximated as going to zero and eq 13 reduces to

$$K_{os} = (4/3)\pi a^3 N_A \times 10^{-3} = a^3 \times 2.52 \times 10^{21} \quad (14)$$

For the ferriporphyrin–imidazole association, an estimate of 9 Å for the center-to-center distance would be reasonable and K_{os} would be ~10. This is a lower limit on the association since it seems likely that a highly charged ion, ferriporphyrin, +7, would have some effect on a molecule with a significant dipole, imidazole (5.7D in naphthalene).[13] An upper limit of approximately 100 would still require that k_{-H_2O} be 9×10^4 sec^{-1}, a thousandfold increase in the rate of water loss for iron(III).[14] The reactivity of imidazole over imidazolium with a positively charged porphyrin can be compared to a similar study of Hasinoff, Dunford, and Horne with negatively charged ferriproto-

(12) D. B. Rorabacher, *Inorg. Chem.*, 5, 1891 (1966).
(13) W. Huckel, J. Datow, and E. Simersbach, *Z. Phys. Chem., Abt. A*, 186, 129 (1940).
(14) See ref 8, p 1042.

porphyrin IX in ethanol-water solutions.[4] They found with ferriprotoporphyrin IX that imidazolium was much more reactive than imidazole. It appears that by changing the charge on the peripheral side chains significant effects can be produced on the metal substitution reactions.

A comparison of the rate constants k'_{12} and k'_{34} shows that the rates of formation of the two imidazole complexes are approximately equal. This is consistent with the proposed mechanism if one assumes that the statistical difference of a factor of 2 in the loss of water from the diaquo- and monaquomonohydroxy complexes is compensated for by an increased rate of water loss from the monohydroxy complex.

The hydrolysis behavior of iron(III) is modified considerably when the iron is complexed to a porphyrin as evidenced by a change in the first hydrolysis constant from $10^{-2.2}$ to $10^{-4.8}$.[15] This occurs despite the fact that charge neutralization should be more important in the case of the positively charged ferriporphyrin.

The use of the ethylenediamine-substituted protoporphyrin IX has thus enabled a more extensive study of the hydrolysis behavior of an iron(III) porphyrin and has given valuable insight into some of the effects of side chains on the substitution behavior of metalloporphyrins. Indeed, the effect of the charge on the side chains has been to give specificity to the heme in reactions with charged and uncharged substrates.

(15) L. G. Sillen and A. E. Martell, *Chem. Soc., Spec. Publ.*, No. 17, 53 (1964).

An Ion Cyclotron Resonance Study of the Structure of $C_3H_6^+$ and the Mechanism of Its Reaction with Ammonia

Michael L. Gross

Contribution from the Department of Chemistry, University of Nebraska-Lincoln, Lincoln, Nebraska 68508.
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Abstract: The reactions of ammonia and ammonia-d_3 and the C_3H_6 radical cation, produced in the ionization of cyclopropane and in the unimolecular fragmentation of tetrahydrofuran, and the $C_3H_4D_2$ radical cation from tetrahydrofuran-2,2,5,5-d_4 were investigated in detail. The $C_3H_6^+$ ion reacts with ND_3, and the elements of C_2H_5 are lost as C_2H_4D indicating no scrambling of hydrogen among the carbon and nitrogen atoms of the collision complex. Furthermore, the label distributions in the product ions produced by $C_3H_4D_2^+$ and NH_3 or ND_3 demonstrate that the structure of $C_3H_4D_2^+$ is one in which the three carbon atoms have become equivalent. The results also show that there is essentially no randomization of hydrogen atoms in the primary $C_3H_4D_2^+$ or in the intermediate complex formed in the reaction with ammonia.

To aid in the interpretation of mass spectra of cyclic hydrocarbons, it has been suggested that ionization removes an electron from a C-C bond, effectively opening the ring.[1] This ring scission is followed by subsequent fragmentation. Supporting evidence for this hypothesis is found in labeling experiments of cyclopentane[2] and methylcyclopentane[3] and in energetic measurements for the fragmentation of methylcyclopentane.[4] The nondecomposing or stable molecular ions from cyclobutane have been shown to be open structures by photoionization followed by product analysis[5] and by ion–molecule reaction studies.[6] In fact, a more recent comparison of the bimolecular reactivities of the molecular ions of cyclobutane, 1-butene, 2-butene, and 2-methylpropene indicates that two or more different open-ring structures occur in the ionization of cyclobutane.[7] Radiolysis of cyclopropane seems to produce a propene ion;[8] however, recent studies of the ion-molecule reactions of $C_3H_6^+$ from cyclopropane indicate that it does not possess this structure.[9,10] These latter results do not exclude the possibility of an unisomerized acyclic ion, *i.e.*, the trimethylene radical cation.

The purpose of the research reported here is twofold. By employing suitably labeled reagents in the previously reported reaction of $C_3H_6^+$ with ammonia (reaction 1),[10] we hope to be able to answer whether the

$$C_3H_6^+ + NH_3 \rightarrow [C_3H_6NH_3^+]^* \begin{cases} \xrightarrow{-C_2H_4} CH_5N \ (m/e\ 31) \\ \xrightarrow{-C_2H_5} CH_4N \ (m/e\ 30) \end{cases} \quad (1)$$

first member of the saturated cyclic hydrocarbon ion series is better represented as an acyclic ion. A second, but equally important, purpose is to test whether randomization of hydrogen atoms occurs in the intermediate complex of reaction 1. Such randomization is

(1) See, for example, J. H. Beynon, R. A. Saunders, and A. E. Williams, "The Mass Spectra of Organic Molecules," Elsevier, Amsterdam, 1968, p 109.
(2) B. J. Millard and D. F. Shaw, *J. Chem. Soc. B*, 664 (1966).
(3) D. P. Stevenson, *J. Amer. Chem. Soc.*, **80**, 1571 (1958); S. Meyerson, T. D. Nevitt, and P. N. Rylander, *Advan. Mass Spectrom.*, **2**, 313 (1963).
(4) R. F. Pottie, A. G. Harrison, and F. P. Lossing, *J. Amer. Chem. Soc.*, **83**, 3204 (1961).
(5) R. D. Doepker and P. Ausloos, *J. Chem. Phys.*, **43**, 3814 (1965).
(6) B. M. Hughes and T. O. Tiernan, *ibid.*, **51**, 4373 (1969).
(7) L. W. Sieck, S. K. Searles, and P. Ausloos, *J. Amer. Chem. Soc.*, **91**, 7627 (1969).
(8) P. Ausloos and S. G. Lias, *J. Chem. Phys.*, **43**, 127 (1965).
(9) L. W. Sieck and J. H. Futrell, *ibid.*, **45**, 560 (1966).
(10) M. L. Gross and F. W. McLafferty, *J. Amer. Chem. Soc.*, **93**, 1267 (1971).