Coordination Chemistry Reviews, 61 (1985) 55–96
Elsevier Science Publishers B.V., Amsterdam – Printed in The Netherlands

KINETICS AND MECHANISMS OF METALLOPORPHYRIN REACTIONS

DAVID K. LAVALLEE

Chemistry Department, Hunter College of CUNY, 695 Park Avenue, New York, New York 10021 (U.S.A.)

(Received 3 April 1984)

CONTENTS

ABBREVIATIONS

H_2P, a free base porphyrin; P, the corresponding porphyrin ligand; MP, the corresponding metalloporphyrin; H_2TPP, 5,10,15,20-tetraphenylporphine; *N*-CH_3HTPP, *N*-methyl-5,10,15,20-tetraphenylporphine; H_2TpyP, 5,10,15,20-tetra-4-pyridylporphine; H_2Ppy_3P, 5-phenyl-10,15,20-triphenyl-

0010-8545/85/$14.70

porphine; $H_2TpyPH_4^{4+}$, 5,10,15,20-tetra-4-pyridiniumporphyrin cation; H_2TMpyP^{4+}, 5,10,15,20-tetrakis(*N*-methyl-4-pyridyl)porphyrin cation (with $H_2TM(2)pyP^{4+}$ and $H_2TM(3)pyP^{4+}$ the corresponding 2-pyridyl and 3-pyridyl isomers, respectively); $H_2TPPS_4^{4-}$, 5,10,15,20-tetrakis(*p*-sulfophenyl)porphinato anion; H_2TAP^{4+}, 5,10,15,20-tetrakis(*N*,*N*,*N*-trimethyl-4-anilium)porphyrin cation; $H_2TM(4)QP^{4+}$, 5,10,15,20-tetrakis(*N*-methyl-4-quinolyl)porphyrin cation (and $H_2TM(3)QP^{4+}$, the corresponding 3-quinolyl isomer); H_2OEP, 2,3,7,8,12,13,17,18-octaethylporphine; H_2DP, deuteroporphyrin (and H_2DPDME, the corresponding dimethyl ester); H_2MP, mesoporphyrin and H_2PP protoporphyrin.

A. INTRODUCTION

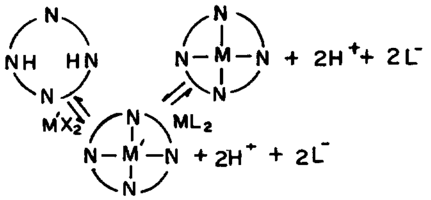
The coordination chemistry of porphyrins has provided a fertile ground for kinetic investigations. The complexation and dissociation reactions of metalloporphyrins have been the principal focus of several reviews, including those of Hambright [1], Schneider [2] and Longo [3], all of which summarize literature to 1974. Berezin's book [4] covers Russian work to 1977 with some coverage of other workers to 1973. The intervening decade has seen many developments in the synthesis and structural chemistry of metalloporphyrins and much has been learned about their biochemical roles. In addition, the photochemistry of porphyrins and their potential for catalysis have brought about productive investigations. Some of these developments have shed light on previous kinetics studies and others have led to new kinetics initiatives. Even an area which has been active for many years, complexation reactions, has been significantly affected by recent developments. A number of observations which were puzzling or seemed contradictory can now be compatibly rationalized.

This review begins with a discussion of complexation and metal atom displacement reactions, including transmetallation and hydrolytic reactions. In the sections dealing with ligation and oxidation–reduction reactions, the specific cases of kinetics and mechanisms of reactions involving oxygen, superoxide and peroxide are discussed. The section on photochemistry includes both ligation and redox reactions. The final sections cover studies of specialized reactions, dealkylation of *N*-substituted metalloporphyrins and metalloporphyrin aggregation.

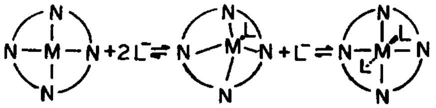
B. METALLATION, TRANSMETALLATION AND HYDROLYSIS

The overall scheme for complexation and protolytic demetallation reactions of metalloporphyrins is diagrammed for a dipositive metal atom in Scheme 1.

Scheme 1



Scheme 2



The metalloporphyrin product can be 4, 5 or 6 coordinate as illustrated in Scheme 2. Some cases of atypical structures have been discussed by Tsutsui [5], but all of the kinetic studies discussed herein have involved the common geometries shown in Scheme 2. To put the discussion of specific studies into perspective, the overall reaction sequence that best accommodates currently published data is presented below (using a dipositive metal ion and neutral ligands and omitting axial ligation equilibria for convenience):

Deformation

$$H_2P \rightleftharpoons H_2P^* \quad K_D \tag{1}$$

Outer-sphere complexation

$$ML_6^{2+} + H_2P^* \rightleftharpoons [L_6M, H_2P^*]^{2+} \quad K_{os} \tag{2}$$

Ligand dissociation and first bond formation

$$[ML_6M, H_2P^*]^{2+} \rightleftharpoons L_5M\text{–}H_2P^{2+} + L \quad K_1 \tag{3}$$

Second bond formation (SAT)

$$L_5M\text{–}H_2P^{2+} \overset{k_{2X}}{\rightleftharpoons} L_nM{=}H_2P^{2+} + (5-n)L \tag{4}$$

Unassisted metalloporphyrin formation

$$L_nM{=}H_2P \underset{k_d}{\overset{k_f}{\rightleftharpoons}} MP + L_n + 2\,H^+ \tag{5}$$

or assisted metalloporphyrin formation

$$MP + 2\,H^+ + M'L_m \underset{k_{d,M'}}{\overset{k_{d,M}}{\rightleftharpoons}} M'P + 2\,H^+ + M^{2+} + mL \tag{6}$$

As discussed below, the slowest step in the sequence is not always the same.

(i) Metallation

Deformation

The only study directly related to the deformation equilibrium is that of Pasternack et al. [6]. They used the temperature-jump method to study the relaxation of the disturbed equilibrium of 5,10,15,20-tetrakis(*N*-methyl-4-pyridyl)porphine (H_2TMPyP) in aqueous solution (see Fig. 1 for the porphyrin numbering scheme). The first-order process they observed at 25°C by monitoring visible absorption changes was assigned to deformation with a forward rate constant of 4.6×10^7 s^{-1} and a reverse rate constant of 1.2×10^9 s^{-1}, giving an equilibrium constant of 0.04. A less direct indication of the deformation requirement is provided by comparison of the ratios of rate constants for reactions of *N*-methyl-5,10,15,20-tetraphenylporphine (*N*-CH_3HTPP) and 5,10,15,20-tetraphenylporphine (H_2TPP) under comparable conditions, which are ca. 10^4–10^5 for Cu(II), Zn(II) and Cd(II) [7,8], and the ratios of rates for the incorporation of Zn(II) and Cd(II) into *N*-methyletioporphyrin and etioporphyrin which are ca. 10^5 [9]. In the case of tetraphenylporphyrins, the basicities of the unmethylated and methylated porphyrins are similar (within 1–2 pH units depending on the solvents used [10,11]). The pK_a difference for methylated and unmethylated porphyrins without *meso* aryl substituents, such as etioporphyrin, is greater [12]. The similar increase in the metallation rates observed for the two cases cited point to distortion rather then increased basicity as the major determinant. The present data are, of course, very limited and could be unrepresentative.

Outer-sphere complexation

Estimates of the outer-sphere complexation constant, K_{os}, are typically

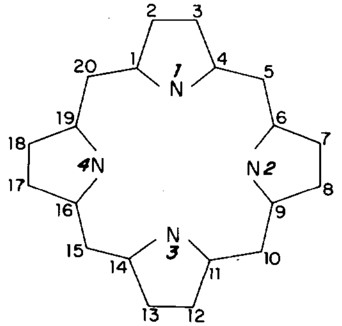


Fig. 1. The numbering scheme for peripheral substituents. Numbering refers to the ligand (e.g. ...*N*-methyl-4-pyridyl...) and is related to the point of attachment of the ligand to the porphyrin's periphery.

made using the Fuoss approach [13], equation 7,

$$K_{os} = 4\pi N_0 a^3 e^b/3000 = 2.54 \times 10^{-3} a^3 e^b \qquad (7)$$

where $4\pi a^3/3$ = the volume of the cation in Å^3, $b = e^2 Z_A Z_B/aDk_BT$ and all other symbols have their usual meanings. In aqueous solution at low ionic strength, K_{os} values have been estimated to be about 10^2 for a dipositive metal complex and $H_2TPPS_4^{4-}$ and about 10^{-2} for a dipositive metal complex and H_2TMpyP^{4+} [2], but the values of K_{os} approach about 3 for both cases at high ionic strength. The data for the reactions of Zn^{2+} with H_2TMPyP^{4+} ($k = 0.52$ M^{-1} s^{-1} at 25°C in 1.0 M $NaNO_3$) and with $H_2TPPS_4^{4-}$ ($k = 0.15$ M^{-1} s^{-1} at 25°C in 1.0 M $NaClO_4$) [2,14,15] are consistent with the similarity of K_{os} values at high ionic strength. As pointed out by Schneider [2], specific ion effects are more likely to affect rates at high ionic strength than are general salt effects. The data of Kassner and Wang (see ref. 2) for reactions of Fe(II) with hemato-, copro- and uroporphyrins (k values of 0.22 M^{-1} s^{-1}, 8.2 M^{-1} s^{-1} and 84.0 M^{-1} s^{-1} respectively, at pH 8.2, $I \leqslant 0.001$ M, 25°C) show the large variation of K_{os} at low ionic strength. Pasternack et al. reported a value for K_{os} in nonaqueous solution [16]. They achieved a very good fit of the curved plot of k_{obsd} vs. $[Cu^{2+}]$ for the reaction of Cu^{2+} with H_2TPP in DMSO using eq. 8

$$k_{obsd} = kK_{os}[Cu^{2+}]/(1 + K_{os}[Cu^{2+}]) \qquad (8)$$

The value they calculate is $K_{os} = 5.5$ at 25°C and $I = 0.2$ M. The most common situations for metallation studies have been aqueous solutions at high ionic strengths and solutions in less polar media (e.g. DMF or DMSO) containing uncharged porphyrins. In either of these cases only small differences in K_{os} for a series of porphyrins or metal ions would be expected.

If the ion-pair or ion-molecular complex is formed at a diffusion-controlled rate and the K_{os} value is near unity, the lifetime of the associated species is less than that required for the formation of the deformed porphyrin species observed by Pasternack et al. [6], i.e. 4.6×10^{-7} s. Thus, to produce a metalloporphyrin, the metal-ion-containing species must, in general, form an outer-sphere complex with the deformed form of the porphyrin. Both the deformation and the outer-sphere complexation reactions are much faster than the overall rate of metalloporphyrin formation and enter the rate expression as pre-equilibrium factors.

Ligand dissociation and first bond formation

Comparative studies of metalloporphyrin formation have been reported for the metal atoms Cu(II), Zn(II), Ni(II), Co(II) and Mn(II) [6,7,14,17]. The typical mechanism for ligand substitution for these ions is presumed to occur via dissociative interchange except for Mn(II), which shows a negative $\Delta V^{\neq}$

TABLE 1

(a) Comparison of rate constants for complexation with tetrakis(4-*N*-methylpyridyl)porphyrin and *N*-methyltetraphenylporphyrin

Ion	Tetrakis(4-*N*-methylpyridyl)porphyrin in H_2O [a]				*N*-Methyltetraphenylporphyrin in DMF [b]			
	k (20°C) ($M^{-1}\,s^{-1}$)	Normal ized rate constant	k_d [a] (H_2O) (s^{-1})	$10^9 k/k_d$, (M^{-1})	k (25°C), ($M^{-1}\,s^{-1}$)	Normal-ized rate constant	k_d [b] (DMF) (s^{-1})	$10^6 k/k_d$ (M^{-1})
Cu(II)	2.3 ± 0.2	46000	2.5×10^9	0.86	289 ± 50	9.6×10^5	2.5×10^8	1.2
Zn(II)	$(50 \pm 3) \times 10^{-3}$	1000	3.2×10^7	1.58	10.4 ± 0.8	3.5×10^4	5.6×10^6	1.8
Co(II)	$(22 \pm 2) \times 10^{-4}$	43	2.5×10^6	0.86	0.68×0.04	2.3×10^3	3.9×10^5	1.7
Mn(II)	$(25 \pm 2) \times 10^{-4}$	50	3.2×10^7	0.08	$0.010 \pm 0.002 (0.14)$ [c]	33	2.4×10^6	0.004(0.06) [c]
Ni(II)	5×10^{-5}	1	2.8×10^4	1.78	0.0003	1	3.8×10^3	0.079

(b) Comparison of the activation parameters for the reaction of the metal with *N*-methyltetraphenylporphyrin and tetraphenylporphyrin in dimethylformamide

Ion	*N*-Methyltetraphenylporphyrin [d]			Tetraphenylporphyrin [e]		
	$\Delta G^{\neq}$ (kcal mol^{-1})	$\Delta H^{\neq}$ (kcal mol^{-1})	$\Delta S^{\neq}$ (eu)	$\Delta G^{\neq}$ (kcal mol^{-1})	$\Delta H^{\neq}$ (kcal mol^{-1})	$\Delta G^{\neq}$ (eu)
Cu(II)	14.0 ± 3.6	16.9 ± 1.8	9.7 ± 6.0	20.4	16.3	-9.8
Zn(II)	16.2 ± 1.5	14.2 ± 0.7	-6.7 ± 2.6	21.6	15.3	-18.3
Co(II)	17.8 ± 2.2	20.4 ± 1.1	8.6 ± 3.6	23.1	24.9	5.2
Mn(II)	20.3 ± 0.9	21.6 ± 0.5	4.5 ± 1.6			
Ni(II)	22.4 ± 3.6	21.4 ± 2.0	-3.2 ± 5.8	23.2	23.4	0.6

[a] Ref. 14. [b] Ref. 8. [c] For the nitrate salts, ref. 21. [d] These determinations were done with metal perchlorate salts, ref. 8. [e] These determinations were done with metal chloride salts, ref. 22.

[18–21]. Recently Tanaka and co-workers [21] reported values for the volumes of activation of formation of the Mn(II), Co(II), Ni(II), Zn(II), Cd(II) and Hg(II) complexes of *N*-methyltetraphenylporphyrin (in cm^3 mol^{-1}, 12.9 ± 0.8, 8.0 ± 0.3, 9.0 ± 0.9, 7.0 ± 0.6, 8.9 ± 1.6 and 4.3 ± 0.8, respectively) which indicate dissociative character for all metalloporphyrin formation reactions.

The results of reactions of H_2TMPyP^{4+} and N-CH_3HTPP in which the first coordination sphere about the metal ion consists of solvent molecules are summarized in Table 1. As first pointed out by Hambright [1], there is a striking correlation of the overall rate constants for metalloporphyrin formation and the dissociation rate of the solvent. The exceptions to the trend are the reactions of Mn(II). It is quite possible that the anomalous behaviour of Mn(II) is due to the associative nature of the reaction in which the first metal-to-porphyrin bond is formed and the reversibility of this step in the reaction sequence. The lack of ligand field stabilization energy for high spin Mn(II) would be expected to lead to a less favorable k_1/k_{-1} ratio than those of other transition metal ions. If the first bond formation is a pre-equilibrium factor in the overall rate expression, that step could be associative even if the overall reaction is dissociative, as indicated by Tanaka's activation volume data. With the Mn(II) data in Table I excluded, the ratio of k_{obs}/k_{ex} for the other metal ions reduces the range of k_{obs} values of 10^6 to a range of only about 10^2.

A number of other studies report metal ion reactivity. Longo et al. have reported rate constants and activation parameters for a number of chloro complexes reacting with H_2TPP in DMF [22]. Although there may be some difficulty in interpreting the data because there are several metal complexes possible in such solutions, the order of the rate constants, Cu(II) > Cd(II) > Zn(II) > Pd(II) > Co(II) > Ni(II) ≫ Mg(II) and the isokinetic relationship of the activation parameters is reasonable in light of the studies discussed above. Longo et al. reported a second order rate law, first order in metal ion concentration as well as porphyrin concentration. Using hexakis(dimethylformamide) complexes of Cu(II) and Zn(II) in DMF, Tanaka and coworkers [7] studied reactions over a wider concentration range than Longo et al. and concluded that a more complex rate expression (see below) was necessary to interpret the data. Tanaka's data match Longo's at high [Cu(II)] and Cu(II) was more reactive than Zn(II) over the entire concentration range.

Berezin et al. studied the reactions of TPP and porphine with acetate salts of several metal ions in unbuffered ethanol [23]. In a previous study of reactions of protoporphyrin and pheophytins, the metal ion dependences of the incorporation reactions were not strictly first-order [24] and in this case second-order rate constants are given without data as a function of metal ion concentration. For 25°C, the calculated rate constants (M^{-1} s^{-1}) for reactions with H_2TPP are 2.05 (Cu(II)), 1.14 (Zn(II)), 1.62 (Cd(II)), 0.051 (Co(II)) and 9.7×10^{-4} (Mn(II)) and those for reactions with H_2P are 0.55 (Cu(II)), 0.22 (Zn(II)), 0.19 (Cd(II)), 5.8×10^{-3} (Co(II)) and 2.7×10^{-4} (Mo(II)).

Choi, Fleischer and coworkers [25,26], Brisbin and coworkers [27–29], and Tanaka and coworkers [30] have studied incorporation reactions for several metal ions in acetic acid or mixed solvents containing acetic acid. Rate laws varied from half-order to first-order in metal ion concentration. In these systems, equilibria involving acetate complexes (as discussed by James and Hambright with respect to Cu(II) equilibria [31]) complicate the interpretation of the effects due to the *d* orbital occupancy and/or size of the metal ion.

Another factor which could affect the rate of bond formation is the nature of the porphyrin ring. The value of the equilibrium constant for formation of the monoprotonated porphyrin pK_3, has been correlated to the rate of formation of water soluble porphyrin complexes of Cu(II) by Adeyamo et al. as shown in Fig. 2 [32]. Some scatter is expected in such a plot since some of the charged substituents which make porphyrins water soluble are relatively close to the complexation site. The results they report for the tetrakis(*N*-methyl-4-quinolyl)- and tetrakis(*N*-methyl-3-quinolyl)porphyrins, in which the *N*-methyl groups are all relatively far from the center of the porphyrin, clearly show the role of basicity. Among uncharged non-*meso*-substituted porphyrin esters and *p*-substituted-*meso*-tetraphenylporphyrins, the pK_3 values vary about 3 units, with electron donating substituents causing a greater basicity for both series [2]. Porphyrin basicity (pK_3) can often be correlated

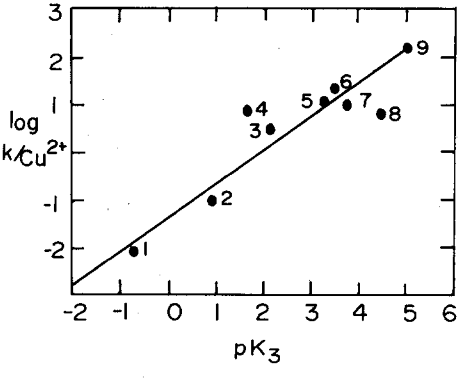


Fig. 2. A plot of $k_0/[Cu^{2+}]$ versus the porphyrin point protonation constant, pK_3 for formation of Cu(II) porphyrin complexes (ref. 32): 1, $H_2TM(2)pyP$; 2, $H_2TM(4)QP$; 3, $H_2TM(4)pyP$; 4, $H_2TM(3)pyP$; 5, $H_2TM(3)QP$; 6, H_2PMpy_3P; 7, H_2TAP; 8, deuteroporphyrin-2,4-disulfonic acid; 9, H_2TPPS_4.

with reduction potentials [33,34]. These potentials are more easily determined experimentally and may prove to be a useful guide to relative reactivities. Even though studies in which Cu(II) complexes have been formed with non-*meso*-substituted porphyrins and pheophytins in ethanol [35] and *p*-substituted-*meso*-tetraphenylporphyrins in DMF [22] show a correlation of rate with basicity, the results are not definitive because both series of reactions show isokinetic behavior and the data were obtained near the isokinetic temperature [1].

A very interesting example of the effect of porphyrin structure on rate is the relatively rapid formation of complexes of Cu(II), Zn(II), Co(II) and Ni(II) with 5,10,15,20-$\alpha,\alpha,\alpha,\alpha$-tetra(*o*-*N*-maleimidoanilio)porphyrin reported by Buckingham et al. [36]. A pre-equilibrium, in which the carboxylate groups bind the metal ion, provides a very favorable configuration for the rate determining step of metal ion transfer to the nitrogen atoms. Turay and Hambright [37] have studied the incorporation of Cu(II) into a related series of porphyrins, the isomers of the picket fence porphyrin, 5,10,15,20-tetra(*o*-pivalamidophenyl)porphine and compared the rates with other *meso*-tetraarylporphyrin reactions. As is evident from Table 2, the bulky pivalamidophenyl groups cause a 100-fold decrease in rate even for the $\alpha,\alpha,\alpha,\alpha$-atropisomer in which one side of the porphyrin is stereochemically similar to the non-*ortho*-substituted species. Reasons given for the slower reactions of a related water soluble picket fence porphyrin include the possibilities of lower basicities, that the *o*-substituted species are less easily deformed, or that proton dissociation may be less favorable because of restriction of solvent from the reaction site [37,38].

Another novel porphyrin, tetra(benzo-15-crown-5)porphyrin, has been used for the study of Zn(II) complexation [39]. The four benzo-15-crown-5 moieties which are found at the periphery can, of course, bind alkali metal

TABLE 2

Rate data for copper incorporation [a]

Porphyrin	$k(s^{-1})$
T($\alpha,\alpha,\alpha,\alpha$-*o*-($(CH_3)_3$CONH)PP	4.6×10^{-4}
T($\alpha,\alpha,\alpha,\beta$-*o*-($(CH_3)_3$CONH)PP	3.8×10^{-4}
T($\alpha,\alpha,\beta,\beta$-*o*-($(CH_3)_3$CONH)PP	3.0×10^{-4}
T($\alpha,\beta,\alpha,\beta$-*o*-($(CH_3)_3$CONH)PP	3.4×10^{-4}
T(*o*-NH_2)PP	7.0×10^{-2}
T(*p*-$(CH_3)_3$CONH)PP	6.0×10^{-2}
T(*p*-(CH_3CONH))PP	5.0×10^{-2}
T(*p*-H)PP	2.6×10^{-2}
T(*p*-CN)PP	1.0×10^{-2}

[a] 25°C, DMF, data from ref. 37.

ions. In 1 : 1 $CHCl_3 : CH_3OH$ solution, this species binds Zn(II) about 30% slower in the presence of 0.1 M $NaClO_4$ than in the presence of tert-butylammonium perchlorate. It binds Zn(II) about 2.5 times faster than does H_2TPP under the same conditions. The effect is quite small and could be due to substituent effects common to the *p*-substituted tetraphenylporphyrins.

Second bond formation and the sitting-atop complex

One of the most important stimuli to the study of metalloporphyrin formation mechanisms was the proposal by Fleischer and Wang of the existence of an intermediate in which the metal atom forms two bonds to the porphyrin while the two hydrogen atoms on the core nitrogen atoms remain bound [40]. Although the existence of isolable sitting-atop (SAT) complexes is not as prevalent as first supposed, it is very likely that such species are common intermediates in the metalloporphyrin formation sequence. The spectroscopic features of the original sitting-atop species were later shown to be similar to ion pairs [41] and fluorescence spectra established that the metal ions were not covalently bound under the reported reaction conditions [42]. Reports of sitting-atop complexes in reactions of 5,10,15,20-tetrapyridylporphine reactions [26] were later shown to be due to salt effects [43]. However, two examples have been presented which strongly support the formation of a sitting-atop intermediate: the intermediate in the reaction of Pd(II) with hematoporphyrin reported by Theophanides and coworkers [44–46] and the intermediate in the reaction of $[Rh(CO)_2Cl]_2$ with H_2TPP reported by Fleischer and Dixon [47]. Grigg et al. have shown that both monometallic and bimetallic products can be formed in such reactions [48]. In both of these cases, the metal atoms are typically stable in square coordination environments. This feature is not characteristic of reactions of transition metal ions of the first row. However, the existence of such species supports the proposition that SAT's are unisolable intermediates in other cases.

Perhaps the strongest case for the common existence of a sitting-atop intermediate is the fact that the reactive form of the porphyrin (typically deduced from kinetics studies as a function of pH) is the free base [1,2,49]. Porphyrins are such weak acids that the deprotonated HP^- form is rare and protonated forms, H_3P^+ and H_4P^{2+}, are unreactive. (The results of a number of pH dependence studies are difficult to interpret, however, because equilibria involving the coordination sphere of the metal ion in addition to those of the porphyrin are affected by pH.)

Unassisted metalloporphyrin formation

In this step, the SAT collapses to form the metalloporphyrin with concerted loss of the hydrogen atoms and insertion of the metal ion. The important question is whether formation of the first bond, formation of the –

SAT (the second bond) or collapse of the SAT is typically rate determining. Current evidence favors formation of the SAT as the slow step. One of the most striking features of metalloporphyrin formation reactions is their extreme sluggishness in comparison with nonmacrocyclic nitrogen donor ligands. The distortion equilibrium constant of 10^{-2} [6] could not account for this retardation unless the amount of distortion observed in Pasternack and coworkers' study is insufficient for bond formation. In this regard, however, it should be noted that the structure of a free base *N*-methyltetraarylporphyrin, *N*-methyl-5,10,15,20-tetra(*p*-bromophenyl)porphine and metal complexes of *N*-methyltetraphenylporphyrin [50,51] are distorted to very similar extents and yet the complexation reactions are still not as fast as those of linear nitrogenous ligands [8]. The anomalous behavior of Mn(II) previously cited is almost most consistent with reversible formation of the first metal–porphyrin bond. The rate laws for hydrolysis of metalloporphyrins, which will be discussed in further detail later, typically indicate that cleavage of the final metal–nitrogen bond is much more rapid than cleavage of the second metal–nitrogen bond. If a single metal–nitrogen bond is readily cleaved, a reversible equilibrium is reasonable.

The question of whether second-bond (SAT) formation or SAT collapse and proton release is rate determining has been investigated by testing for kinetic isotope effects. In metallation reactions of *N*-CH_3HTPP in DMF, Bain-Ackerman and Lavallee showed that there was no effect due to addition of D_2O [8]. Lavallee and Onady [52] have interpreted the observed kinetic isotope effect for the formation of ZnTPP from H_2TPP or D_2TPP in dimethylformamide solutions with added water or deuterium oxide in terms of changes in the first coordination sphere of Zn(II) rather than an effect of the hydrogen or deuterium atoms on the porphyrin nitrogen atoms. These results indicate that SAT formation rather than proton release is rate determining.

Assisted metalloporphyrin formation

In a study of the reactions of Zn(II) with CdTPP, PbTPP and HgTPP in pyridine, Grant and Hambright found that Zn(II) could react with a metalloporphyrin, HgTPP, to form ZnTPP at a rate faster than the direct reaction of Zn(II) with H_2TPP [53]. The addition of Hg(II) increased the rate of conversion of PbTPP to ZnTPP by Zn(II). Hambright, Tanaka and coworkers have shown that the formation of manganese porphyrins in aqueous solution from Mn(II) salts is greatly accelerated by addition of Pb(II), Cd(II) or Hg(II) salts [54–57]. The fact that these transmetallation reactions are actually faster than the reaction of the free base eliminates the free base as an intermediate. Since the porphyrin used by Tanaka, $H_2TPPS_4^{4-}$, is a very weak acid, a free anionic form such as $HTPPS_4^{5-}$ or $TPPS_4^{6-}$ would be protonated to form the free base at diffusion-controlled rates. Because this process would generate the free base in a highly favorable pre-equilibrium step, it can be eliminated as a mechanism. In addition, the spectral overlaps presented by Grant and Hambright [53] for the Zn(II) + HgTPP reaction show seven well-defined isosbestic points, supporting the proposition of a direct reaction of Zn(II) with the intact HgTPP molecule.

Tanaka has also proposed the presence of homodinuclear species in the reactions of cobalt(II) acetate with H_2TPP [58] and solvated Cu(II) with H_2TPP in DMF [7] in order to explain the non-linear dependence of the rate on metal-ion concentration.

Thus, one metal atom can bond to a free base porphyrin and thereby accelerate the formation of a complex by a second metal atom. All of the distorting metal atoms which have been reported form stable four-coordinate complexes and can be readily dissociated from their porphyrin complexes by acid. In addition, they either are known to form out-of-plane complexes or, where structures are not known, they are so large that such distortion is expected [59,60]. In these cases, then, it may be relatively easy for metal–porphyrin bonds to be broken and reformed. It seems reasonable that if a pyrrole ring could then rotate, exposing a lone pair, a second metal ion could bind. Further bond-breaking, ring rotation and bond formation can lead to the new metalloporphyrin.

Assisted metallation through a metalloporphyrin intermediate is only expected for weakly-bonding metal ions that have out-of-plane structures. The reactions are expected to lead to the product with the strongest metal–nitrogen bonds. Catalysis occurs in cases in which the weaker-binding metal ion binds more rapidly (i.e. in the case of acceleration of porphinatomanganese(II) formation, the Cd(II) or Hg(II) complexes are formed rapidly and when these ions are displaced, formation of the manganese complex is irreversible).

Other means of catalyzing metalloporphyrin formation have included the use of microemulsion media, in which rate enhancement is proposed to be due to an alteration in the equilibria of the precursor metal complexes [61], and addition of ligands (e.g. acetate, tironate and 8-oxyquinolate, [2,62]) which have reasonably high equilibrium constants but fast dissociation rates. Whitten and coworkers investigated Cu(II) incorporation into a series of porphyrine containing different alkyl substituents using monolayer films and assemblies as well as DMF solutions [63]. The hexadecyltetraamine derivative of $\alpha,\alpha,\alpha,\alpha$-tetraaminophenylporphyrin gave the most rapid reaction in assemblies, indicating that it is probably present at the hydrophilic interphase. The reactivities of the various derivatives were quite different, showing the importance of the microenvironment.

Summary

The slow step in the formation of metalloporphyrins is typically the formation of an SAT intermediate (with the special exception of transmetallation reactions). The forward rate of this step will be predominantly determined by desolvation of the metal ion to allow formation of the second bond (i.e. a dissociative interchange step), as shown by the correlation of overall formation rates with solvent exchange rates [7,8,14] and volume of activation data [19]. The overall mechanism is relatively complex, however, and the overall rate constant and the rate law can readily reflect the variation in pre-equilibrium constants as conditions are changed. The effect of peripheral substituents on the formation rate is relatively small and favorable outer-sphere complexation is only pronounced for oppositely charged species at low ionic strength.

(ii) Transmetallation

One type of transmetallation reaction, catalytic transmetallation, has already been mentioned as an example of assisted metalloporphyrin formation. Several studies have also been made of other reactions in which an isolated metalloporphyrin is dissolved and a salt added to cause transmetallation.

The quadridentate, relatively inflexible macrocyclic porphyrin ring generally exerts a pronounced kinetic chelate effect, making it resistant to metal atom displacement. Studies of homonuclear metal exchange have involved radioactive isotopes or, in other cases, solutions of a metalloporphyrin and a free base porphyrin in which distinct spectral change would accompany metal atom exchange [64–68]. Exchange was demonstrated for sodium complexes [65], which are highly ionic and relatively labile. Exchange is also expected in the presence of small amounts of acid for metalloporphyrins which are sensitive to protolytic demetallation. Iron porphyrins undergo exchange in resorcinol melts [69].

Exchange has been observed for complexes of metal ions which typically form tetrahedral complexes with unidentate ligands and form out-of-plane metalloporphyrins (e.g. Zn(II), Cd(II), Pd(II) and Hg(II)) when the displacing ion forms a more stable metalloporphyrin [53–64]. The most comprehensive study has been reported for replacement of Cd(II) by Zn(II) in a wide variety of porphyrins (Tables 3 and 4; Fig. 3) [70]. An isokinetic relation was found (isokinetic temperature = −30°C), as has been shown previously for the incorporation of $Cu(OAc)_2$ into various porphyrins in ethanol ($T_i = 22°C$) [31] and reactions of $CuCl_2$ in DMF ($T_i = 57°C$) [22]. In the high temperature region, electron donating groups on the periphery (which increase basicity) increase the rate of exchange. It is typical that the

TABLE 3

Rate data for *meso*-substituted cadmium–porphyrin exchange with zinc

Cd–Porphyrin [a]	$\Delta H^{\neq}$ (kcal mol^{-1})	$\Delta S^{\neq}$ (eu)	$k_{25°C}$	k_{exch}/k_{incor} [b] (M^{-1} s^{-1})
(p-CH_3)-TPP	16.8	+2.1	8.3	1640
(o-CH_3)-TPP	14.6	−5.2	8.5	6700
(p-OCH_3)-TPP	13.9	−7.3	10.2	1000
(p-CN)-TPP	13.1	−14.0	1.2	1400
(p-H)-TPP	11.5	16.4	5.7	1530
(p-Cl)-TPP	10.8	−21.6	1.3	600

[a] Pyridine–0.5 M water solvent, with $Zn(ClO_4)_2$, ref. 69. [b] k_{exch}/k_{incor} is the ratio of the Zn/Cd–P exchange rate to the Zn/H_2–P incorporation rate at 25°C.

TABLE 4

Rate data for β-substituted cadmium–porphyrin exchange reactions with zinc, 25°C [a]

Cd–Porphyrin [b]	$k_{Zn/Cd-P}$ (M^{-1} s^{-1})
Etioporphyrin-I	41.8
Octaethylporphyrin	37.6
Coproporphyrin-III-TME	20.2
Uroporphyrin-I-OME	17.4
Protoporphyrin-IX-DME	14.2
Deuteroporphyrin-IX-DME	13.8
Br_2Deuteroporphyrin-IX-DME	3.4
Diacetyldeuteroporphyrin-IX-DME	3.1

[a] Pyridine–0.5 M water solvent, with $Zn(ClO_4)_2$, ref. 69. [b] OME, TME, DME are the octa-, tetra-, and di-methylesters.

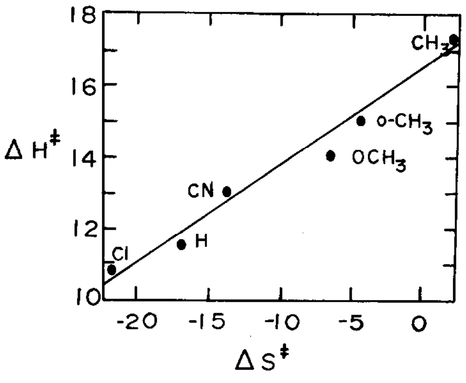


Fig. 3. A plot of $\Delta H^{\neq}$ vs. $\Delta S^{\neq}$ for conversion of CdP to ZnP with P = a 5,10,15,20-tetrakis(*para*-substituted phenyl)porphyrin except for the case of o-CH_3. The reaction involved $Zn(ClO_4)_2$ dissolved in pyridine containing 0.5 M water (ref. 70).

value of $\Delta H^{\neq}$, which is the determining factor above the isokinetic temperature, is lower for more basic porphyrins not only for transmetallation but also for complexation [35] and protolytic demetallation [23]. In aqueous ammoniacal solutions, Zn(II) or Co(II) replace the Mn(II) atom of a tetrakis(ethylenediamine) derivative of PPIX [71]. Porphyrin complexes of Mn(II) are reactive in a variety of aqueous media, including ammoniacal solutions, but the exchange reaction for Zn(II) is faster than the direct metallation by Zn(II) under the same conditions (at pH 7.4, $[NH_4^+] = 1.09$ M, $[NH_3] = 9.2 \times 10^{-3}$ M, 25 M^{-1} s^{-1} and 1.7 M^{-1} s^{-1}, respectively), indicating that the free base porphyrin is not an intermediate.

Stinson and Hambright have shown that Cd(II) can be displaced from its *N*-methyltetraphenylporphyrin complex by Cu(II) [72]. In this case, the Cu(II) atom cannot attack from the side opposite the out-of-plane Cd(II) atom and the authors invoked a same-side direct electrophilic substitution mechanism. The data show the reaction to be first-order with respect to [Cu(II)] with no evident suppression upon addition of Cd(II) [72]. Subsequently, it was demonstrated by Lavallee et al. that the reaction is also first order with respect to the concentration of water [68]. That result and preliminary data for the acid hydrolysis of Cd $N\text{-}CH_3TPP^+$ lead to suggest an alternate mechanism that does not involve "same-side" attack of the incoming metal ion. The hydrolysis of *N*-methyltetraphenylporphinatocadmium(II) follows the rate expression typical of other *N*-methylporphyrin complexes (vide infra)

$$k_{obs} = [H^+]^2/(a + b[H^+]) \tag{9}$$

For the system containing Cd $N\text{-}CH_3TPP^+$ and Cu(II), hydrolysis of Cd $N\text{-}CH_3TPP$ can result from $[H^+]$ produced by the cupric ion

$$Cu^{2+} + H_2O \rightleftharpoons CuOH^+ + H^+ \quad K_1$$

In this case, the $[H^+]$ is such that $a \gg b[H^+]$. If Cu(II), rather than directly displacing Cd(II), reacts with $N\text{-}CH_3HTPP$ in a relatively fast step, the expected rate law would be

$$\text{Rate} = [H^+]^2[\text{Cd}N\text{-}CH_3TPP]/a = (K_1/a)[Cu^{2+}][H_2O][\text{Cd}N\text{-}CH_3TPP]$$

consistent with the experimental result [68] (if $[Cu^{2+}] \sim [Cu]_{tot}$)

$$\text{Rate} = k[Cu^{2+}][H_2O][\text{Cd}N\text{-}CH_3TPP] \tag{10}$$

(iii) Protolytic demetallation

Metalloporphyrin stability is often defined in terms of the degree of resistance toward displacement of the metal atom by acid [73]. This definition is used because it is seldom possible to determine equilibrium constants directly (although some values have been determined kinetically for moderately stable complexes such as those of Zn(II) and Cd(II) [74,75], Table 5) and because metalloporphyrins show a wide range of sensitivity to acid. Some metalloporphyrins in which the bonding is likely to be highly ionic—alkali metal and alkaline earth porphyrins—undergo demetallation so readily that the reactions are difficult to study. Several filled *d* shell complexes, those of Hg(II), Pd(II), Cd(II) and Zn(II), are demetallated by dilute acid solutions. It is also possible to remove Mn(II) and Fe(II) readily. Hence, most of the available kinetic data apply to reactions of Zn(II), Cd(II), Mn(II) and Fe(II) porphyrins.

Protolytic demetallation of zinc complexes of two synthetic porphyrins, $ZnTPPS_4^{4-}$ and $ZnTMPyP^{4+}$, has been studied in aqueous solution. The reaction of $ZnTPPS_4^{4-}$ with acid ($HClO_4$, $I = 0.1$ M using $NaClO_4$, 30°C) gave the rate law:

$$\text{Rate} = k_d[H^+]^2[\text{Zn–TPPS}_4^{4-}] \qquad (11)$$

over an acid concentration of 1 to 43 millimolar, with $k_d = 11\ M^{-2}\ s^{-1}$ [75]. The observed pseudo-first-order rate constant was insensitive to the use of sodium halides in place of $NaClO_4$. With the positively charged ZnTMPyP, however, the demetallation reaction is very sensitive to the nature of the counterion [76,77]. At constant ionic strength

$$\text{Rate} = k_d[\text{ZnTMpyP}][H^+]^2 \qquad (12)$$

Such acid solvolysis (as well as metal incorporation and exchange reactions) shows typical Bronsted–Bjerrum ionic strength behavior. When extrapolated to $\mu = 0$, at 25°C, $k_d = 86\ M^{-2}\ s^{-1}$ for Zn_4^{4-} and $6.0 \times 10^{-5}\ M^{-1}\ s^{-1}$ for $ZnTMpyP^{4+}$ [78].

Shamim and Hambright reported acid hydrolysis rates for Cd(II) com-

TABLE 5

Equilibrium and rate data for cadmium porphyrins (25°C)

Porphyrin [a]	pK_3 [b]	K_e^{Cd} [c,d]	k_f ($M^{-1}\ s^{-1}$)	k_r ($M^{-2}\ s^{-1}$)
H_2TM(2)pyP	0.9	7.9×10^{-7}	6.9	8.8×10^{6}
H_2TM(4)pyP	1.4	2.0×10^{-8}	39.6	2.0×10^{9}
H_2TM(3)pyP	1.8	7.7×10^{-9}	51.7	6.7×10^{9}
H_2PM-py$_3$P	2.4	6.1×10^{-9}	91.2	1.5×10^{10}
H_2TAP	3.6	1.7×10^{-10}	86.6	5.1×10^{11}

[a] Ref. 70. [b] pK_3 is the porphyrin monocation dissociation reaction (H_3P^+/H_2P). [c] The ionic strength is 0.2($NaNO_3$). [d] $K_e^{Cd} = (CdP)H^+)^2/(H_2P)(Cd^{2+})$, where $K_e^{Cd} = k_f/k_r$.

plexes of several porphyrins: the 2, 3, and 4 isomers of tetrakis(N-methylpyridyl)porphyrin, 5-phenyl-10,15,20-tris(N-methyl-4-pyridyl)porphyrin($H_2PM\text{-}py_3P$) and tetrakis(N,N,N-trimethyl-4-anilium)-porphyrin (H_2TAP) [74]. All of these species show second-order proton dependence in 0.2 M $NaNO_3$ at 25°C

$$\text{Rate} = k_d[\text{MP}][\text{H}^+]^2 \tag{13}$$

with results shown in Table 5. The acid dissociation rate constants are directly related to the basicity of the free base porphyrin corresponding to each complex (Fig. 4). The relative rates have also been correlated with Büchler's stability index [73], as shown in Fig. 5 [79]. The dependence of k_d on pK_3 is much stronger than the dependence of metalloporphyrin formation rate, suggesting the importance of the stability of the protonated intermediate formed prior to metal atom dissociation. Second-order proton dependence has also been found for Fe(II) and Mg(II) porphyrins [80].

Hambright and coworkers have also studied reactions of tetrakis(N,N,N-trimethyl-4-anilium)porphinato complexes of Cu(II), Co(II), Ni(II) and In(III) [79–81]. The observed rate law is:

$$\text{Rate} = k_d[\text{M–TAP}](h_0)^n \tag{14}$$

where h_0 is the Hammett acidity function. In the studies of the reactions of Cu(II), Co(II) and Ni(II), the n values were found to be above 2 but the possible role of the counterion was difficult to assess because of the instability or low solubility of the porphyrins and the high sensitivity of the

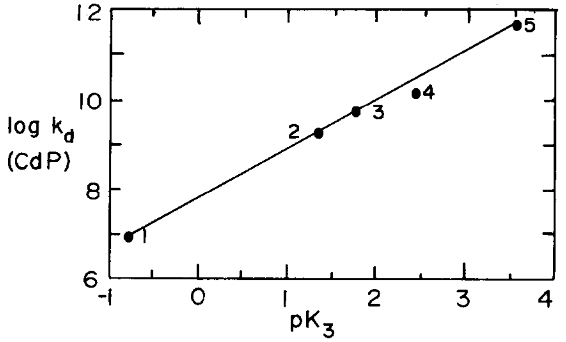


Fig. 4. A plot of the log of the CdP dissociation rate constants, k_d(CdP), vs. the porphyrin protonation constant, pK_3 (ref. 49). 1, $H_2TM(2)pyP$; 2, $H_2TM(4)pyP$; 3, $H_2TM(3)pyP$; 4, $H_2PM\text{-}py_3P$; 5, H_2TAP.

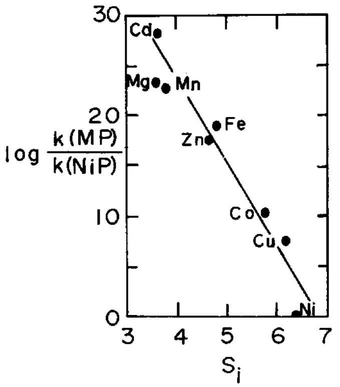


Fig. 5. A plot of the log of the relative rates of hydrolysis of MP and the corresponding nickel(II) porphyrin, log(K_{MP}/k_{NiP}), vs. the stability index S_i (ref. 80).

rate to small differences in acid concentration [79,80]. For the In(III) complexes of H_2TMPyP^{4+}, H_2TAP^{4+} and $H_2TPPS_4^{4-}$, however, the picture is much clearer [81]. For the rate law

$$\text{Rate} = k_d[\text{In}^{3+}\text{-P}][h_o]^n \tag{15}$$

at 25°C, $n = 2.2$ and $k = 1.1 \times 10^{-6}$ M^{-n} s^{-1} for $InTMPyP^{5+}$ in HCl; $n = 2.4$ and $k_d = 8.7 \times 10^{-6}$ M^{-n} s^{-1} for $InTAP^{5+}$ in HCl; $n = 2.4$ and $k_d = 9.3 \times 10^{-6}$ for $InTPPS_4^{3-}$ in HCl and $n = 2.1$ and $k_d = 2.3 \times 10^{-6}$ for $InTPPS_4^{3-}$ in HNO_3. $InTPP_4^{3-}$ is stable for long periods in $HClO_4/NaClO_4$ and dissociation is greatly catalyzed by Cl^- or SCN^- as is the reaction of $ZnTMPyP^{4+}$ [74,75], in contrast to the lack of dependence found for the reaction of $ZnTPPS_4^{-4}$ [76]. In the study of the acid solvolysis of $FeTPPS_4^{3-}$ in mixed acid–water–ethanol media by Reynolds et al. [83] and Hambright [84], the demetallation rate was found to vary with the medium in the order $HCl > H_2SO_4 > HClO_4$ [83,84]. The similarities of the solution chemistries of Fe(III) and In(III) are thus reflected to some extent in the chemistries of their porphyrin complexes.

Hambright and coworkers have reported acidic demetallation kinetics for Mn(II) complexes of isomeric tetra(*N*-methylpyridyl)porphyrins [85,86]. The *N*-methyl-2-pyridylporphyrin is the most acidic porphyrin known ($pK_3 < 0$, [86]) and the Mn(II) complex has a half-life in 1.0 M HCl (extrapolated) of about 38 ms compared with 74 μs for the Mn(II) complex of the more basic *N*-methyl-3-pyridylporphyrin. In addition to the difference in rate, these isomeric complexes show different rate laws. The pseudo-first-order rate constant for reaction of the 2-pyridyl complex has the form

$$k_0 = k_A[H^+] + k_B[H^+]^2 \tag{16}$$

with $k_A = 1.8 + 0.1$ M^{-1} s^{-1} and $k_B = 16.4 + 0.6$ M^{-2} s^{-1} at 25°C, $\mu = 0.5$ (NaCl/HCl) [85] whereas the 3-*N*-pyridyl complex has an observed rate constant of the form:

$$[H^+]/k_0 = a[H^+]^{-1} + b \tag{17}$$

with $a = (7.2 \pm 0.6) \times 10^{-6}$ and $b = (10.2 \pm 1.2) \times 10^{-5}$ at 25°C, $\mu = 0.5$ (NaCl/HCl) [86]. The proposed mechanism

$$\text{Mn(II)P} + H^+ \rightleftharpoons [\text{Mn(II)P–H}^+]$$

$$[\text{Mn(II)P–H}^+] + H^+ \rightarrow \text{products}$$

with $[\text{Mn(II)P–H}^+]$ as a steady-state intermediate gives:

$$k_0 = k_1k_2[H^+]^2/(k_{-1} + k_2[H^+]) \tag{18}$$

Then $k_1 = b^{-1} = (1.0 \pm 0.1) \times 10^4$ M^{-1} s^{-1} and $k_{-1}/k_2 = a/b = (7.2 \pm 0.8)$

$\times 10^{-2}$ M [86]. As noted by Hambright, this is the only example of competition for a protonated intermediate in aqueous solution but it is consistent with the behavior of hematoporphinatomanganese(II) in H_2O/EtOH [87].

A few demetallation reactions have been studied in mixed aqueous–organic media. In water–propane–toluene microemulsion media, the hydrolysis of chlorophyll a is second order in acid concentration. In microemulsions of the composition χ(toluene) = 0.411, $\chi(H_2O) = 0.180$ and χ(2-propanol) = 0.408 at 25°C and $I = 0.10$ M ($NaClO_4$), for chlorophyll *a*, $k_d = 4.7 \times 10^3$ M^{-2} s^{-1} [88]. For the same conditions, the rate law for demetallation of bacteriochlorophyll a is first order with $k_d = 1.1$ M^{-1} s^{-1} while the rate law for MgTPP is mixed first and second order, $k_0 = k_A[H^+] + k_B[H^+]^2$, with $k_A = 0.04$ M^{-1} s^{-1} and $k_B = 24.4$ M^{-2} s^{-1} [89]. The half-lives at $[H^+] = 10^{-2}$ M are 0.3 s for chl a, 13 s for bchl a and 4 min for MgTPP. The reaction rates are highly sensitive to the composition of the microemulsion media, with increasing rates as the proportion of hydrocarbon is increased [90]. The reaction order is also a function of the medium. In 2-propanol, the demetallation of MgTPP is third order in acid with $k = 83$ M^{-2} s^{-1} while the reaction of bchl a is still first order in acid, with $k = 0.15$ M^{-1} s^{-1} [88]. At $[H^+] = 10^{-3}$ m, in 2-propanol the half-lives are very similar, 3 min for bchl a, 4 min for chl a and 8 min for MgTPP. We interpret the differences in the behavior of these complexes as the result of the weaker basicity and metal–nitrogen bond strength of the saturated pyrrole moieties of chl a (with one) and bchl a (with 2). The metal–nitrogen bonds to the saturated rings may be rapidly cleaved after the rate determining step, leading to a lower order rate law.

Acid solvolysis in alcohols has also been reported for Mg(II) complexes [91]. Using methanol with $HClO_4$, Snellgrove and Plane [91] found the rate law for the demetallation of deuteroporphyrin IX dimethyl

$$\text{Rate} = k[\text{MgP}][\text{H}^+]^3/(\rho + [\text{H}^+]) \tag{19}$$

where $k = 3.1 \times 10^5$ M^{-2} s^{-1} and $\rho = 5 \times 10^{-5}$ M at 42°C. In ethanol under the same conditions the same rate law obtains, with $k = 1.2 \times 10^3$ M^{-2} s^{-1} and $\rho = 4 \times 10^{-5}$ M. Thus, the rate is very sensitive to the medium. In the presence of pyridine (MeOH–pyridine–$HClO_4$ medium), the rate law includes pyridine

$$\text{Rate} = k[\text{MgP}][\text{py}][\text{H}^+]^3/(\rho + [\text{H}^+]) \tag{20}$$

with $k = 7.2 \times 10^8$ M^{-3} s^{-1} and $\rho = 5 \times 10^{-6}$ M at 42°C. With water added, only second-order dependence on acid was observed

$$\text{Rate} = k[\text{MgP}][\text{py}][\text{H}_2\text{O}][\text{H}^+]^2 \tag{21}$$

with $k = 2.7 \times 10^9\ \mathrm{M}^{-4}\ \mathrm{s}^{-1}$ at 42°C.

The acid solvolysis of Zn(II) complexes in methanol has been reported for etioporphyrin and *N*-methyletioporphyrin complexes [92]. The rate law for the former is

$$\mathrm{Rate} = k[\mathrm{ZnP}][\mathrm{H}^+]^3/(\rho + [\mathrm{H}^+]) \tag{22}$$

with $k = 2.0 \times 10^5\ \mathrm{M}^{-1}\ \mathrm{s}^{-1}$ and $\rho = 6.6 \times 10^{-3}$ M at 25°C, while for the latter,

$$\mathrm{Rate} = k'[\mathrm{ZnN\text{-}CH_3P}][\mathrm{H}^+]^2/(\rho' + [\mathrm{H}^+]) \tag{23}$$

with $k' = 2.5 \times 10^2$ M and $\rho' = 0.25$ M at 25°C. The lower order very likely reflects the fact that one of the four zinc–nitrogen bonds in the *N*-methylporphyrin complex is weak [51] so that the zinc atom can readily dissociate after two rather than three nitrogen atoms have been protonated. As noted previously, we have found the same rate law for the demetallation of $(\mathrm{Cd}N\mathrm{CH_3TPP})^+$. Although it is recognized that *N*-alkylated porphyrins are typically demetallated at quite different rates (under the same conditions) than corresponding non-*N*-alkylated porphyrins, there are no other kinetic studies for quantitative comparison.

In the same report [92], Hambright and coworkers also showed the hydrolysis rates for Zn(II) complexes of non-*meso*-substituted porphyrins parallel the free ligand basicity ($\mathrm{p}K_3$), with more basic porphyrins losing the Zn(II) ion more rapidly at a given $[\mathrm{H}^+]$ (k_1 in s^{-1} and $\mathrm{p}K_3$ values are: 1.59 and 5.85 for mesoporphinatozinc(II); 0.57 and 5.5 for deuteroporphinatozinc(II); 0.130 and 4.8 for protoporphinatozinc(II) and 1.30×10^{-3} and 3.0 for dibromodeuteroporphinatozinc(II); all as dimethyl esters with $[\mathrm{HCl}] = 7.5 \times 10^{-3}$ M in methanol at 25°C).

Berezin and his coworkers have published numerous reports of protolytic demetallation reactions of metalloporphyrins as well as pheophytin and phthalocyanine complexes [4,93]. The favored solvents are media such as t-butanol and trichloroacetic acid, DMSO and sulfuric acid, ethanol and acetic acid, glacial acetic acid, or glacial acetic acid and sulfuric acid. In such media, determination of hydrogen ion concentration is problematic. The general findings are that the stability order is Fe(III) > Ni(II) > Cu(II) > Zn(II) > Mg(II) > Cd(II) and that mesoporphyrin complexes undergo dissociation more readily than tetraporphyrin complexes [4,93–97]. They have found that Mn(III), Al(III) and Sn(IV) complexes of TPP are very resistant to protolytic demetallation [94].

Nickel and vanadyl porphyrins present in crude oil can cause problems with catalysts used in the refining processes. These complexes are very difficult to demetallate. Hung and Wei have developed improved catalytic hydrodemetallation processes using high pressure $\mathrm{H_2}$ (up to 12500 kPa), a

$CoO-MoO_3-M_2O_3$ catalyst and an oil solvent medium [98]. After the porphyrin ring is hydrogenated, the ring is fragmented and the metal ion is deposited on the catalyst. Under the same conditions, metal-free porphyrins rapidly decompose. Such methods appear to be more feasible than acidic solvolysis.

In some cases demetallation has been accomplished by first reducing the metalloporphyrin to give a metal in a lower oxidation state that is more easily removed. This method has been applied to Fe(III) porphyrins (in which HCl and Fe(II) act as cooperative catalysts [99]) as well as Ag(II) and Ag(III) porphyrins [100,101]. Oxidative demetallation has been achieved with $W(V)TPPS_4^-$ [102] and Kadish and Riffard have reported an interesting demetallation reaction that is presumably nonacidic [103]. In this reaction, mercury is removed from HgTPP by imidazole in two steps, a relatively rapid one involving formation of Hg(TPP)(Im) followed by the slow step which is first order in imidazole and Hg(TPP)(Im).

The substantial literature on demetallation reactions shows several general features: (1) since the highest order of acid concentration is third, it appears that a single metal–nitrogen bond can be readily broken, (2) the rates and even the rate law are very sensitive to the solvent, (3) strongly complexing anions can accelerate reactions, (4) more basic porphyrins can be solvolytically demetallated more rapidly and (5) demetallation rates are inversely dependent on metalloporphyrin stability (as defined by Büchler [73]).

C. LIGAND EXCHANGE REACTIONS

Hambright [1] and Büchler [73] have discussed the thermodynamics of axial ligation. This section will specifically concern kinetic and mechanistic studies.

Axial ligand exchange can be more rapid for metalloporphyrin complexes than for complexes of the same metal ion with all monodentate ligands, as first shown by Fleischer and coworkers [104]. For Co(III) porphyrins, several research groups have reported labilization factors of about 10^6 [104–111]. Rate enhancements have also been found for saturated macrocyclic complexes of Co(III) [112]. These were interpreted in terms of differences in Co–N bond lengths and release of strain energy in forming a five coordinate intermediate and could not be accounted for in terms of *cis* effects of the macrocycles [112,113]. The Co–N bond lengths of Co(III) porphyrins, however, are in the normal range (1.96–1.99 Å [108]). However, Williams and Hambright have shown that *cis* effects are probably important since the rates of thiocyanate anation of a series of Co(III) porphyrins are directly related to the basicity of the porphyrin free ligand with $\log k_f = 0.5\ pK_3 - 0.3$ for pK_3 values varying by over 4 units [114]. The mechanism proposed for ligand substitution for the Co(III) porphyrins is dissociative.

The labilization factor for Cr(III) porphyrins has been a subject of some controversy [116–120], with estimates of 0 to 10^3, but labilization has been clearly shown for complexes of LCr(III)(TPP)Cl where L is a variety of PR_3 or $P(OR)_3$ species or pyridine in toluene [120]. In this medium, complications due to solvent coordination and deprotonation are avoided. The proposed mechanism is dissociative and formation of Cr(III)(TPP)Cl intermediate is strongly supported [120]. For substitution of water by imidazole, Krishnamurthy has found that *trans* OH^- is labilizing and that neither the imidazolium ion nor the anion of imidazole replace H_2O [117]. Leipoldt et al. [119] also found that OH^- is labilizing; anation of Cr(III)(TMPyP) $(H_2O)(OH)^{4+}$ by CN^- or NCS^- about 5×10^3 faster than of the diaquo complex, an effect similar to that found previously for Cr(III)$TPPS_4$ complexes [118]. Studies have also been carried out on ligand exchange on another typically inert ion, Rh(III). Krishnamurthy has reported that the replacement of H_2O by SCN^- or CN^- on Rh(III)$TPPS_4$ occurs at a comparable rate to H_2O replacement on $[RhCl_5H_2O]^{2-}$ [117]. The mechanism in this case is not as clear as for the Co(III) porphyrins, but the data for the reaction of CN^- are consistent with an ion-pair associative interchange mechanism.

Ligand exchange on Fe(III) is generally faster for porphyrin complexes than for $Fe(H_2O)_6^{3+}$. In a study of replacement of water by imidazole on 6,7-bis(ethylenediamine)protoporphyrin IX dimethyl ester iron(III), Kolski and Plane [121] proposed an ion-pair exchange mechanism

$$(H_2O)_2FeP^+ + Im \rightleftharpoons [Im\ldots(H_2O)_2FeP]^+ \rightleftharpoons Im(H_2O)FeP^+ + H_2O$$

such that the overall formation constant K (3.8×10^3) and rate constants ($k_f = 8.5 \times 10^6\ M^{-1}\ s^{-1}$ and $k_r = 1.8 \times 10^3\ s^{-1}$) can be used to estimate a k_d value (from $k_d = k_f/K_{os}$ and $K_{os} \sim 100$) of $\sim 9 \times 10^4\ s^{-1}$, about 10^3 faster than water exchange on $Fe(H_2O)_6^{3+}$. La Mar and coworkers have studied the replacement of halide ions on XFeP (where P = *para* substituted tetraphenylporphyrins) by alkylimidazoles (L) to form L_2FeP^+ complexes [122]. Although the individual rate constants for the two ligand replacement steps cannot be deduced from the data, the half-lives for each overall reaction (to form six-coordinate complexes) vary from those typical of Fe(III) substitution reactions (~ 0.1 s for 5-CH_3In replacing Cl^-, Br^- or I^-) to considerably shorter ones (e.g. $\sim 10^{-3}$ s for 4-CH_3py replacing I^-). Rate variations for tetraarylporphyrins with different *p*-phenyl substituents are small (a factor of 3 from *p*-Cl to *p*-CH_3O) but the corresponding complex of the more basic octaethylporphyrin reacts about 10 times faster.

Reactions of imidazole and 1-methylimidazole with iron(III) porphyrins have been studied by Pasternack and coworkers [123,124] and by Thompson and Krishnamurthy [125]. Pasternack found that equilibrium constants for complexation of imidazole or 1-methylimidazole in Me_2SO were similar for $Fe(TPP)^+$ or hemin but that the reactions with hemin were faster and exhibited the relationship $\tau^{-1} = k_f[L]^2 + k_r$ only at low concentrations of ligand indicating a possible mechanistic difference (interpreted as the binding of the second imidazole to hemin before the high-spin to low-spin conversion). Krishnamurthy studied the reactions of the μ-oxo Fe(III) dimers of $TPPS_4$ and TAP with imidazole. The negatively charged $TPPS_4$ dimer reacts about 40 times faster than the TAP derivative at 30°C. From the pH dependence of the reactions, the reactive ligand species is the imidazolium ion. Generally, only the bis(imidazole)Fe(III) complexes are observed, but Burdige and Sweigart have used a bulkier alkylimidazole, *N*-propylimidazole and did find kinetic evidence for the reversible formation of a mono(*N*-propylimidazole) species [126]. Stynes and coworkers have found that exchange of imidazole is significantly slower for Ru(II) porphyrins than Fe(II) porphyrins and that the *cis* effect is much less pronounced for the ruthenium complexes [127,128].

In a most interesting study of axial ligand exchange in Fe(III) porphyrins, Snyder and LaMar used NMR line broadening to demonstrate that iron atom inversion can accompany halide ion exchange [129]. The rate law for exchange is second order (rate $= k[\mathrm{PFeX}][\mathrm{X}^-]$) for exchange of tetrabutylammonium salts in $CDCl_3$. The $\Delta S^{\neq}$ values reported were highly negative ($\sim -20 \pm 3$ e.u.), consistent with an associative mechanism.

The binding of CO to Fe(II) porphyrins having one imidazole axial ligand is of interest with regard to the bonding of CO to hemoglobin. Studies by Hoffman and coworkers [130,131] indicate that π-donor/acceptor interactions with the macrocyclic porphyrin ring do not influence rates of CO bonding to as great an extent as imidazole deprotonation (which causes the binding rate of CO with Fe(TPP)(Im) to drop from 1.1×10^7 M^{-1} s^{-1} to 3.4×10^4 M^{-1} s^{-1}) and it was suggested that hydrogen bonding could have a pronounced effect in the protein. In a study of the rates of CO binding by Fe(II) complexes of the "capped" porphyrins 5,10,15,20-[pyromellitoyl(tetrakis-*o*-oxyethoxyphenyl)]porphyrin and 5,10,15,20-[pyromellitoyl(tetrakis-*o*-oxypropoxyphenyl)]porphyrin and high affinity state hemoglobin, Hoffman and coworkers found comparable rate constants ($\sim 4 \times 10^6$ M^{-1} s^{-1}), showing that CO binding is not sterically restricted by these "caps" [132].

Tanaka and coworkers [133] studied peroxide binding by TiO(TPyP), finding the rate law:

$$\mathrm{d}\left[\mathrm{Ti(O_2)(TpyPH_4)^{4+}}\right]/\mathrm{d}t = k\left[\mathrm{TiO(TpyPH_4)^{4+}}\right]\left[\mathrm{H_2O_2}\right]\left[\mathrm{H^+}\right] \qquad (24)$$

with $k = 55$ M^{-2} s^{-1} at 25°C and $\Delta H^{\neq} = 8.3$ kcal mol^{-1}, $\Delta S^{\neq} = 23$ e.u. and

$\Delta V^{\neq} = -3.3$ cm^3 mol^{-1}, providing evidence for an associative mechanism.

Caldin and Field have used an improved microwave T-jump method to measure very fast ligand binding, finding that nitrogen bases such as imidazole and pyridine bind to ZnTPP in aprotic solvents such as toluene and chlorobenzene at nearly diffusion-controlled rates ($\sim 10^{10}$ M^{-1} s^{-1}) [134].

D. OXIDATION–REDUCTION REACTIONS

The roles of metalloporphyrin prosthetic groups in electron transport proteins and metallochlorins in photosynthesis have led to numerous attempts to determine factors that influence the rates and mechanisms of metalloporphyrin electron transfer reactions. The relative rigidity of the macrocyclic ring and its extensive delocalization tend to minimize the structural changes due to electron transfer and thereby minimize the Franck–Condon contribution to the activation energy. In addition, the participation of the metal atom and the entire ring system in the molecular orbitals involved in electron transfer make it possible for rapid electron transfer that involves the metal center to be initiated at the periphery of the ring. Sutin [135], Gray and coworkers [136], Castro and coworkers [137] and others have investigated reactions of metalloproteins such as cytochrome c with inorganic and organic redox reagents; this work has principally concerned the modifications of metalloporphyrin reactivity (including electrostatic effects [138]) by the surrounding protein, and is not the subject of this review. An aspect which is relevant to studies of the metalloporphyrins themselves, however, is the typical finding that the kinetics of redox reactions of the cytochromes with small molecules are consistent with outer-sphere electron transfer. This mechanism is consistent with general structural features of cytochromes, i.e., an exposed heme edge and relatively restricted access to the metal atom.

(i) Electrochemical studies

Kadish and coworkers have determined electron transfer rates using cyclic voltammetry [139,140]. The peak separations are a function of potential scan rate and are different for peaks assigned to metal-centered and ring-centered processes, indicating faster heterogeneous electron transfer to the ring. The rates are enhanced by electron-donating substituents and more basic porphyrin rings. Davis and coworkers have reported that the basicity of nitrogenous axial ligands and the ability of polar solvents to stabilize the transition state by charge compensation or by altering axial ligand binding constants (e.g. by favoring imidazole vs. chloride ion coordination) enhance heterogeneous electron transfer rates of iron porphyrins [141,142]. Rates for heterogeneous electron transfer reactions of Fe(III) porphyrins had been thought to be sensitive to spin state [e.g., 143,144], but Kadish has recently shown that the Fe(III/II) rate constants for $Fe(OEP)(3\text{-}Clpy)_2^+$ which is predominantly high spin and for $Fe(OEP)(4\text{-}NH_2py)_2^+$ which is predominantly low spin as well as for $Fe(TPP)L_2^+$ complexes are all about 2×10^{-2} $cm^2\ s^{-1}$ [145]. The differences observed previously are attributed to changes in solvation or axial ligation. Kadish and coworkers have also found that electron transfer rates of manganese porphyrins can be orders of magnitude slower than those of iron porphyrins, a difference attributed to the larger structural changes induced by reduction of Mn(III) porphyrins [146].

(ii) Homogeneous reaction kinetics

Pasternack, Heineman, and coworkers have used both cyclic voltammetry spectroelectrochemistry (OTTLE cell) and homogeneous reaction kinetics (with $Ru(NH_3)_6^{2+}$ as reductant) to study reactions of Co(III)TMPyP [147,148]. Their results are in accord with rapid electron transfer through the π system. They calculate a self-exchange rate of about 20 $M^{-1}\ s^{-1}$ which is in close agreement with another self-exchange reaction with little structural change that is presumed to involve electron transfer through a π system, $Co(phen)_3^{3+/2+}$ (21 $M^{-1}\ s^1$). Chapman and Fleischer used NMR spectroscopy to determine the self-exchange rate of $CoTPP(py)_2$ in $CDCl_3$ finding a similar value, 10 $M^{-1}\ s^{-1}$ at 25°C [149]. The exchange rate for the reaction of $Co(III)TPP(py)_2$ with Co(II)TPP(py) is 0.1 $M^{-1}\ s^{-1}$, a retardation that is consistent with the structural difference. Inner-sphere exchange with I^- as bridging ligand is 10^6 faster than the outer-sphere exchange rate [150]. Other ions also catalyze the exchange, in the order $I^- > N_3^- \sim SCN^- > Br^- > Cl^-$.

To determine whether electron transfer occurs axially on through the π system, Fleischer and Cheung reacted $Co(III)TPPS_4$ and tetra(*p*-carboxyphenyl)porphinatocobalt(III) ($Co(III)TPPC_4$) with $^{51}Cr(II)$. They concluded that at least some peripheral reduction occurred because ^{51}Cr eluted with the cobalt porphyrins from cation exchange resin [151]. This conclusion was disputed, however, by Balahura and Trivedi who found that the hydrogen ion dependence ($k_{obsd} = k_0 + k_2[H^+]^{-1}$ where $k_0 = 173 \pm 10$ $M^{-1}\ s^{-1}$ and $k_2 = 2.0 \pm 0.1$ s^{-1} at 25°C, $I = 1.0$ M, $LiClO_4$) and the amount of Cr(III) eluted with $Co(III)TPPS_4$ are consistent with axial reduction [152]. The results of Balahura and Trivedi are consistent with those for the reduction of Co(III)TMPyP by Cr(II) reported by Pasternack and Sutin ($k_0 = 16$ $M^{-1}\ s^{-1}$ and $k_2 = 1.6$ s^{-1} at 25°C [153]). The rates obtained by Pasternack et al. for the reduction of Fe(III)TPyP by Cr(II) [154] are remarkably similar to those for $Co(III)TPPS_4$ and Co(III)TMPyP ($k_0 = 270$

M^{-1} s^{-1} and $k_2 = 8.0$ s^{-1} at 25°C). Reid and Hambright have reported k_2 values for Fe(III)DPDME and Fe(III)MPDME (aqueous 2% triton X-100 solutions, 0.5 M $NaClO_4/HClO_4$, 25°C) of 11.0 s^{-1} and 1.6 s^{-1} [155] (note that their nomenclature differs from ref. 140).

In the reactions of pentacyanocobalt(II) and hexacyanocobalt(II) with a number of dicyano(porphinato)iron(III) complexes, Worthington and Hambright [156] found a rate law of the form: $k_{obs} = [k_1 + k_2(CN^-)][Co^{2+}]$ showing comparable reactivity of the pentacyanocobalt(II) species (possibly inner-sphere) and the hexacyanocobalt(II) species (outer-sphere) as shown in Table 6.

In another study of outer-sphere electron transfer with inorganic reagents, Pasternack and Spiro found that the reduction of Fe(III)TMPyP by $Ru(NH_3)_6^{2+}$ as a function of pH shows that the self-exchange rate of the low-spin hydroxo complexes ($> 10^9$ M^{-1} s^{-1}) is much faster than that of the high-spin aquo complexes (1.2×10^6 M^{-1} s^{-1}) [157]. Based on the results of Kadish and Su for heterogeneous reactions [145], however, this effect could be due to differences in the coordination of the metal ion rather than the spin state per se.

Ong and Castro [158] addressed the question of inner-sphere vs. outer-sphere electron transfer using nitro aromatics to oxidize Fe(II) porphyrins. They found that nitro aromatics react via an inner-sphere mechanism by a two-electron process in which either the first or second electron transfer could be rate determining and toward which six-coordinate low-spin Fe(II) complexes were unreactive. They determined the self-exchange rates for the six coordinate iron porphyrins using NMR line broadening techniques, with rate constants (1.2×10^4 M^{-1} s^{-1} for bis(pyridine-d_5), 1.4×10^5 M^{-1} s^{-1}

TABLE 6

Pentacyanocobaltate(II) reductions of dicyano(porphinato)iron(III) complexes [a]

$(NC)_2Fe^{III}-P$	k_1 (M^{-1} s^{-1}) [b]	k_2 (M^{-2} s^{-1}) [b]	E^{red} (volts) [b]	k (M^{-1} s^{-1}) ($SO_2^-/(NC)_2Fe^{III}-P$)
Mesoporphyrin	40	70	−0.22	8.0×10^7
Deuteroporphyrin	120	340	−0.21	9.0×10^7
Protoporphyrin	80	480	−0.18	1.5×10^8
3,8-Dibromo-deuteroporphyrin	390	390	−0.15	1.8×10^8
meso-Tetra(4-sulfophenyl)porphyrin	400	< 70	−0.14	–
3,8-Diacetyldeuteroporphyrin	72	60	−0.12	2.6×10^8

[a] Ref. 156, unless otherwise specified. [b] 25°C, $I = 0.5$ ($NaOH/NaCN/NaClO_4$), pH = 12.
[c] Reduction potential at pH = 9 for $(NC)_2Fe^{III}-P + e^- = (NC)_2Fe^{II}-P$ (vs. NHE).

for bis(N-methylimidazole), $> 1.4 \times 10^7$ for bis(imidazole) and $> 2 \times 10^7$ for bis(t-butylamine)) that are faster than axial ligand exchange, demonstrating an outer-sphere mechanism [159]. The order of rates for self-exchange parallels those for oxidation of low-spin Fe(II) porphyrins by oxygen [159] and by quinone [160].

Other studies of redox reactions of iron-porphyrins have included reactions with ascorbic acid and alkyl thiols. Toppen and coworkers have reported rate constants for reactions of Fe(III)TMPyP complexes with ascorbic acid [161,162]. In the presence of histidine, the rate law is $k_{\mathrm{obsd}} = k_{\mathrm{diss}}(1 + K_a[\mathrm{H}^+]^{-1})$ where $K_a(1.4 \times 10^{-8}$ M) and $k_{\mathrm{diss}}(25\ \mathrm{s}^{-1})$ are the equilibrium constants for the bis(histidine) complex and the rate constant for histidine dissociation and the reaction is independent of reductant concentration, indicating relatively rapid inner-sphere reduction by bound ascorbate [161]. Without histidine, the rate law is

$$k_{\mathrm{obsd}} = k'_x k'_d [\text{ascorbic acid}] (k'_m + k'_x [\text{ascorbic acid}])^{-1} \qquad (25)$$

where k'_x, k'_d and k'_m represent $[\mathrm{H}^+]$ dependent rate constants for electron transfer, ferriporphyrin dimer formation and ferriporphyrin dimer dissociation, respectively. The rate constant for dimer formation (2 Fe(por)OH → dimer) is $2.1 \times 10^5\ \mathrm{M}^{-1}\ \mathrm{s}^{-1}$ and the rate constant for electron transfer from the ascorbic acid to the ferriporphyrin dimer is $280\ \mathrm{M}^{-1}\ \mathrm{s}^{-1}$ [153]. For these high spin species, electron transfer appears to be a two-electron outer-sphere process [162]. Wayland and Swartz [163] found that Fe(TPP)Cl reacts with alkyl thiols in toluene in the presence of piperidine to form $\mathrm{Fe(TPP)(py)_2}$ and $\mathrm{R_2S_2}$ ($\mathrm{R = CH_3, (CH_3)_3C}$) with the rate law

$$-\mathrm{d}[\mathrm{Fe(TPP)Cl}]/\mathrm{d}t = k[\mathrm{Fe(TPP)Cl}][\mathrm{RSH}][\mathrm{py}]^3 \qquad (26)$$

indicating the involvement of two low-spin intermediates, i.e., Fe(TPP)(py)(SR) and $\mathrm{Fe(TPP)(py)_2(SR)}$.

(iii) Reactions involving oxygen, superoxide and hydrogen peroxide

As mentioned above, Castro et al. deduced an outer-sphere mechanism for the oxidation of six-coordinate Fe(II) porphyrins in amine solvents [160]. From the pH dependence, the rate determining step was proposed to be the dissociation of a protonated iron porphyrin–oxygen π complex. Reactions of a number of Fe(II)TMPyP complexes with $\mathrm{O_2}$ in aqueous solution have been studied by Solomon et al. [164]. The reaction of the aquo complex is first order in $\mathrm{O_2}$ and second order in Fe(II)TMPyP, suggesting the formation of a μ-peroxo-bis(Fe(III)TMPyP) intermediate. Evidence for such an intermediate has been reported by Balch and coworkers from low-temperature NMR spectra [165,166]. Collman, Anson and coworkers have reported a comprehensive study of cofacial metalloporphyrins, finding that species with two cobalt atoms at the appropriate distance for peroxo complex formation were much more active catalysts than species with cobalt atoms at greater distances or heterometallic species [167]. Superoxide dismutase activity has been shown for Fe(III)TMPyP monomer and dimer [168–171], with the monomer being more reactive with a rate constant for O_2^- disproportionation of $\sim 3 \times 10^7\ M^{-1}\ s^{-1}$) and the dimer forming a long-lived intermediate [171]. Solomon et al. [164] found a rate constant of $2 \times 10^9\ M^{-1}\ s^{-1}$ for the reaction of the aquo complex of Fe(III)TMPyP with O_2^- and rate constants 10^3 lower for the dicyano, dihistidyl and bis(imidazole) complexes. The reactions of the six-coördinate low-spin complexes are still much faster than ligand exchange and are outer-sphere. The product of the reaction of O_2^- with the aquo complex of Fe(III)TMPyP is rapidly oxidized by the superoxide radical ($k = 2.3 \times 10^7\ M^{-1}\ s^{-1}$).

Several studies have concerned catalysis of dioxygen or peroxide by adsorbed porphyrins. Anson and coworkers have observed catalytic reduction of oxygen by both Co(II) [172] and Fe(III) [173] porphyrins adsorbed on graphite electrodes. The Co(II) porphyrin also catalyzes both the electroreduction and oxidation of H_2O_2. In the study of the Fe(III) porphyrins, H_2O_2 was detected as an intermediate using a rotating ring disk electrode (RRDE) and alteration in the catalytic behavior of the metalloporphyrin as a result of adsorption were discussed. Bettleheim and Kuwana studied the catalytic electroreduction of O_2 using FeTMPyP in solution using the RRDE and cyclic voltammetry [174]. They also found H_2O_2 (95% of the initial product) and estimated the rate of Fe(II)TMPyP removal by oxygen to be 10^7 to $10^8\ M^{-1}\ s^{-1}$. Additional results obtained from cyclic voltammetry and thin layer coulometry studies of O_2 to H_2O by FeTMPyP support a $2 + 2$ mechanism with initial H_2O_2 binding to form the Fe^{III}–O_2^- intermediate with subsequent reduction by Fe(II)TMPyP and protonation [175].

Van Veen et al. studied oxygen reduction catalysis of a number of metalloporphyrins adsorbed on carbon in strongly acidic (8 N H_2SO_4) and in strongly alkaline (6 N NaOH) media [176]. In acid the activity was correlated with the M(II/III) potential. Mochido et al. found that adsorption of CoTPP on alumina on NiO enhanced its ability to catalyze H_2O_2 decomposition but that a variety of other metal complexes of TPP lost their activity (all exhibited about the same rather small activity at pH 10 when unsupported) [177].

Rocklin and Murray used Co(II) and Cu(II) tetra(*p*-aminophenyl)porphyrins covalently bound to electrodes to catalyze the reduction of dibromoalkyls [178]. Rates obtained using the RRDE for a variety of substrates and different amounts of bound porphyrins indicate that charge diffusion through the porphyrin layer is fast relative to substrate diffusion through the layer and that specific substrate–metalloporphyrin interactions are important in determining the rate.

Very rapid reduction of Zn(II), Co(III), Mn(III) and Al(III) and Sn(IV) porphyrins (at or near the diffusion limit) has been effected by radicals generated via pulse radiolysis [179–182]. Alkyl radicals generated in this manner bond axially to Co(II)$TPPS_4^{4-}$ with rate constants of $(1.2–1.9) \times 10^9$ M^{-1} s^{-1} [180]. Hydroxyradicals which add to Co(II)$TPPS_4^{4-}$ undergo heterolytic cleavage of the Co–C bond to form Co(I)$TPPS_4^{5-}$ and a carbonyl with rate constants of 3.6×10^2 s^{-1} for $R = CH_2OH$ and 6.2×10^3 s^{-1} for $R = C(OH)(CH_3)_2$. Reduction of Mn(III)P to Mn(II)P by e_{aq}^- takes place with $k \sim 10^{10}$ M^{-1} s^{-1} while reduction by radicals from iso-PrOH is about 10 times slower [181]. Hydrolytic demetallation of the Mn(II)P formed by these reactions occurs at rates found for Mn(II)P species produced by conventional reduction reactions [85,86]. When Al(III)TPP^+ and Sn(IV)TPP^{2+} are reduced by pulse radiolysis, the final products are phlorins or chlorins [182].

E. PHOTOCHEMISTRY

The effects of the metal ions in porphyrin photosensitizers and related studies of excited state lifetimes have been the subject of several studies. Advantages of metalloporphyrins for photosensitization include their extremely high absorbance in the visible region, relatively long excited state lifetimes and facile synthetic variation. Metalloporphyrins are often more satisfactory than free-base porphyrins because they are less susceptible to degradation.

Cauzzo et al. found that the rate of photooxidation of methionine is enhanced by hematoporphyrin complexes of Mg^{2+}, Zn^{2+}, Cu^{2+} and Fe^{2+} as well as by the free base [183]. They proposed that the complexes were less active than the free base because of more efficient nonradiative decay of the first excited singlet state and a reduction of the triplet lifetime, consistent with the typical behavior of heavy atoms as quenchers. Rougee et al. used $ZnTPPS_4^{4-}$ to sensitize the oxidation of cysteine (CysSH) by methyl viologen (MV^{2+}). The reaction of the triplet state of photoexisted $ZnTPPS_4^{4-}$(^{3}Por) gives Por^+ (cage yield ~ 50%, with $k = 6.9 \times 10^4$ M^{-1} s^{-1}) which then reacts with a large excess of CysSH ($k = 1.7 \times 10^6$ M^{-1} s^{-1}) forming CyS–SCys which in turn reduces another equivalent of MV^{2+} ($k = 8 \times 10^8$ M^{-1} s^{-1}) [184].

To obtain basic data for the possible application of Zn(II) porphyrins as photosensitizers in solar energy conversion devices, Houlding and coworkers analyzed the effect of ionic strength on triplet decay of ZnTMPyP and on a related zwitterionic complex, tetrakis(*N*-(propylsulfonato)-4-pyridiniumyl)-

porphinatozinc(II) [185]. The rate constants (extrapolated to zero ionic strength, temperature unreported) are 350 s^{-1} and 1990 s^{-1} for unimolecular decay, 1.9×10^7 M^{-1} s^{-1} and 1.5×10^9 M^{-1} s^{-1} for triplet–triplet annihilation and 5×10^6 M^{-1} s^{-1} and $< 10^6$ M^{-1} s^{-1} for ground-state quenching, respectively. The variation in rates with ionic strength was analyzed using simple Debye–Hückel theory with the triplet decay of the zwitterionic species insensitive and that of ZnTMPyP very sensitive. They mentioned that other porphyrins (e.g. H_2TMPyP^{4+} and H_2TAP^{4+}) show similar behavior but the analysis of kinetic data is complicated by ground state aggregation. To the same end, Richoux and Harriman studied the influence of electrostatic forces on the efficiency of charge separation in the Zn porphyrin–methyl viologen system, finding it to be directly related to the charge on the zinc porphyrin [186]. Cationic porphyrins were quenched much more slowly than anionic complexes. Graeztel and coworkers found that the rate of the ZnP* reaction with MV^{2+} in surfactant assemblies is directly related to the concentration of MV in the middle [187]. Back transfer from reduced MV to oxidized ZnTPP could be intercepted by a donor such as NaOH within the micelle. They also found that the back reaction with $ZnTPPS_4^{4-}$ is reduced when MV^{2+} is replaced by neutral *N,N′*-bis(carboxyethyl)-4,4′-bipyridine (CEB) [188]. When $ZnTMpyP^{4+}$ is adsorbed on dihexadecyl phosphate vesicles, decay of the excited state occurs both via triplet ions and by π-cations and π-anions. The latter pathway is enhanced by increasing the surface area, consistent with a concentration mechanism. The rate of oxidative quenching of the triplet by alkyl viologens is similar ($k_Q \sim 10^9$ M^{-1} s^{-1} at 25°C) for $ZnTMpyP^{4+}$ and $ZnTPPS_4^{4-}$ [189]. Potter and Levin used Eu^{3+} as an oxidant and found rates of 4.2×10^5 M^{-1} s^{-1} and 2.1×10^6 M^{-1} s^{-1} for the ZnTPP and MgTPP triplets, respectively [189]. Excitation of Fe(III), Cu(II), Co(III) and V(IV)O complexes did not result in a sufficiently long-lived triplet for similar flash photolysis studies.

In the absence of an oxidant such as methyl viologen, zinc octaethylporphyrin (ZnOEP) undergoes photochemical ionogenesis. Ballard and Mauzerall found that the process occurs in two steps, i.e., elecron transfer in the encounter complex of ground state ZnOEP and excited (triplet) ZnOEP* followed by a collision between two triplets [190]. Neither of the rate constants depends directly on the dielectric constant of the medium.

Zinc(II) complexes have been popular candidates for study because of their relatively high quantum yields, reasonable hydrolytic stability and ease of preparation. Lever and coworkers have found, however, that other metalloporphyrins can catalyze the photoreduction of methyl viologen even more efficiently and that the rates and quantum yields can be predicted to some degree from the driving force of the reaction [191].

Kapinus and Stangi determined the effect of solvents and temperature on the kinetics of triplet state deactivation of diamagnetic species such as Ga(TPP)OH, Ti(TPP)O and In(TPP)OH as well as H_2TPP by organic oxidizing agents. From the rate constants in 29 organic solvents they found some correlation with increasing solvating ability and viscosity but the relationship is not straightforward and they proposed that the interaction of the oxidizing agent with the porphyrin is very important [192]. The temperature dependence of the decay kinetics does not follow the Arrhenius relation, pointing to the importance of several independent factors [193]. Rentzepis and coworkers have used picosecond laser spectroscopy to study the relaxation processes of some paramagnetic metalloporphyrins (with Fe(III), Ag(II) and Cu(II) [194,195]. The Cu(II) complexes have excited state lifetimes (~ 1 μs at room temperature) significantly longer than typical lifetimes of diamagnetic fluorescent porphyrins (in the nanosecond range). The Fe(III) and Ag(II) undergo very fast intersystem crossing and radiationless decay to the ground state. A transient absorption with a lifetime of 31 ps has been attributed to a triplet state of Fe(TPP)Cl [196].

Although most kinetics studies of metalloporphyrin photochemistry have been concerned with sensitization or electron transfer, there have been a few reports involving axial ligand dissociation as the primary photoevent. Substitution of SCN^- on $CoTCPP(H_2O)(SCN)^{4-}$ by photodissociation has been studied by Pasternack and Parr [106]. Six coordinate Fe(II) porphyrins with nitrogenous ligands undergo photodissociation to pentacoordinate species with a quantum yield (~ 0.05) comparable to that of oxyhemoglobin [197]. The recombination reactions are first order with respect to the free ligand in excess and the rate constant is very sensitive to the nature of the ligand. Ledon et al. found that oxomethoxy(5,10,15,20-tetraphenylporphinato)Mo(V) catalyzes the photoassociated reduction of O_2 to H_2O_2 in benzene/methanol solution [198]. The reaction involves cleavage of the Mo–O bond to form the methoxyl radical. The oxidation of the oxomolybdenum(IV) complex by O_2 is second order with respect to Mo(O)TPP, indicating a two-electron process. Although the nature of the intermediate is undetermined, the coordinative unsaturation of the Mo(IV) porphyrin and the rate law suggest a μ-oxygen-bis(Mo(O)TPP) intermediate. Olson et al. have studied the binding and photodissociation of 11 alkyl isocyanides and protoheme mono-3-(1-imidazoyl)propylamide monomethyl ester in benzene and aqueous soap suspensions [199]. The quantum yield (0.5) and the rate constant (0.8–0.9 s^{-1}) for photodissociation were independent of solvent and had little dependence on the size of the ligand. The association rate constants in benzene ($\sim 2 \times 10^8$ s^{-1}) were also unaffected by ligand size but in the soap suspensions the rates decreased from $\sim 2 \times 10^8$ M^{-1} s^{-1} for *n*-butyl isocyanide to $\sim 2 \times 10^7$ M^{-1} s^{-1} for acetonitrile. This decrease was attributed to differences in partition coefficients.

F. DEALKYLATION OF *N*-SUBSTITUTED METALLOPORPHYRINS

In 1970 Shears and Hambright reported that chloro-*N*-methyloctaethylporphinatozinc(II) is converted to ZnOEP upon overnight reflux in pyridine [200]. In 1977, while attempting to study the rates at which metal ions bind to *N*-methyltetra-(5,10,15,20)phenylporphyrin(N-CH_3HTPP), we found that dealkylation occurred in dimethylformamide at room temperature in the presence of Cu(II), with CuTPP being formed. These conditions for cleaving a methyl group from the nitrogen atom of an aromatic nitrogen heterocycle are unusually mild. The possibility of readily forming dealkylated metalloporphyrins was so intriguing that we decided to investigate the reaction in some detail.

A general characteristic of the nucleophilic dealkylation of *N*-substituted metalloporphyrins is a second order rate law (first order in metalloporphyrin and nucleophile concentrations) at low nucleophile concentration. At higher nucleophile concentrations in some situations, notably Cu(II) and Ni(II) complexes in acetonitrile with dialkylamine nucleophiles [201–205], the rate law becomes second order in nucleophile concentration, giving the overall rate law

$$\frac{\mathrm{d}[\mathrm{CuTPP}]}{\mathrm{d}t} = \left\{\frac{k_1 + k_2 K_{\mathrm{eq}}[\mathrm{nu}]}{1 + K_{\mathrm{eq}}[\mathrm{nu}]}\right\}[\mathrm{nu}][\mathrm{Cu}N\text{-}\mathrm{RTPP}^+]$$

where k_1 and k_2 are the rate constants at low and high nucleophile concentrations and K_{eq} is the constant attributed to the equilibrium for bonding of the nucleophile to the *N*-alkylporphyrin complex. There is no independent evidence (e.g. spectroscopic) for the intermediate but the equilibrium constants derived from the kinetics are at least plausible. The path at higher nucleophile concentrations has a smaller rate constant ($k_2 < k_1$) as expected for $K_{eq} > 1$, leading to a downward curvature of the plot of the observed pseudo-first-order rate constant vs. nucleophile concentration, as shown in Fig. 6. The rate constants and activation parameters in this section are those of the k_1 path unless otherwise specified. The rate of these reactions is sensitive to the nature of the nucleophile, with amines more effective than nitrogen heterocyclics and water a poor nucleophile, consistent with the S_N2 character of the reaction. The reaction is slower in highly polar solvents where the nucleophile is better solvated and proceeds via an ion-pair precursor in solvents of low polarity with Cl^- as nucleophile [204], also consistent with S_N2 character.

Of great importance with regard to the rapid synthesis of radiolabelled compounds for medicinal applications [e.g., ref. 207], the rate of dealkylation is highly dependent on the nature of the metal ion (Table 7). This depen-

TABLE 7

Activation parameters for the demethylation of *N*-methyltetraphenylporphyrin complexes by di-n-butylamine in acetonitrile [a,b]

Metal ion	$\Delta H^{\neq}$ (kcal mol^{-1})	$\Delta S^{\neq}$ (eu)	$\Delta G^{\neq}$ (318 K), (kcal mol^{-1})
Pd(II) [c]	14.4 ± 0.6	−21.5 ± 1.6	20.8
Cu(II)	16.9 ± 1.0	−13.1 ± 2.9	21
Ni(II)	18.0 ± 1.0	−14.0 ± 3.0	23
Zn(II)	41.6 ± 2.0	−54.6 ± 6.2	26
Mn(II)	> 51	> 65	

[a] Error limits are the standard deviations from the calculated least-squares fit in (k/T) vs. $1/T$. Eyring activation parameters are specified. [b] Data taken from ref. 202. [c] Ref. 203.

dence is best correlated with the stability of the product metalloporphyrin, favoring those with a planar rather than out-of-plane structure, and suggesting that the activated complex highly resembles the product (i.e., the Hammond postulate). This tendency is exemplified by the fact that high-spin Fe(II) *N*-methylporphyrins are demetallated by nucleophiles such as thiolates rather than dealkylated [206]. This characteristic and the fact that the dealkylation rate is strongly dependent on the nature of the *N*-substituent of the metalloporphyrin, with good carbocation stabilizing substituents most rapidly removed [205] (Table 8), demonstrates the significant S_N1 character of the reaction.

Most of the metal ions we have studied form *N*-alkylporphyrin complexes much more rapidly than they undergo dealkylation. The exceptions revealed so far are Ni(II) and Pd(II). For Pd(II), the rate of complexation can be regulated by use of different solvents. In CH_3CN solutions, there are clear isosbestic points for initial spectra of the Pd(II)N–CH_3TPP complex and

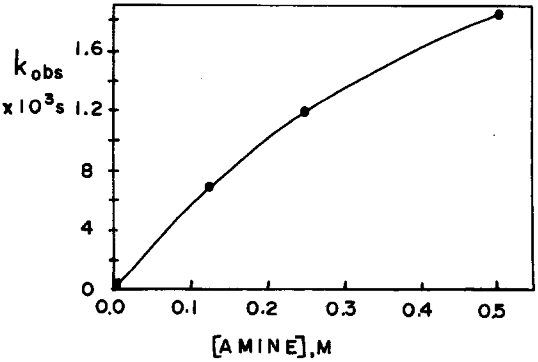


Fig. 6. A plot of the observed pseudo-first-order rate constant for the reaction of *N*-methyldeuteroporphyrin in acetonitrile at 44.2°C vs. the concentration of di-n-butylamine (ref. 204).

TABLE 8

Activation parameters for dealkylation of copper(II) *N*-alkylporphyrin complexes by di-n-butylamine in acetonitrile[a]

Complexes	Path 1[b]			Path 2[c]		
	$\Delta H^{\neq}$ (kcal mol^{-1})	$\Delta S^{\neq}$ (e.u.)	$\Delta G^{\neq}$, 298 K (kcal mol^{-1})	$\Delta H^{\neq}$ (kcal mol^{-1})	$\Delta S^{\neq}$ (e.u.)	$\Delta G^{\neq}$, 298 K (kcal mol^{-1})
$CuN\text{-}CH_2C_6H_4ONO_2TPP^+$	14.8 ± 1.0	-11.6 ± 3.4	18.3 ± 2.0	24.9 ± 2.6	20.3 ± 9.0	18.9 ± 5.3
$CuN\text{-}CH_3TPP^+$	16.9 ± 1.0	-13.1 ± 2.9	20.8 ± 1.9	17.8 ± 0.4	-14.7 ± 1.4	22.2 ± 0.9
$CuN\text{-}C_2H_5TPP^+$	18.2 ± 1.0	-13.9 ± 3.3	22.4 ± 2.0	17.9 ± 5.0	-18.5 ± 16.4	23.4 ± 9.9

[a] Ref. 205. [b] Predominant path at low [di-n-butylamine]. [c] Predominant path at high [di-n-butylamine].

TABLE 9

Activation parameters for dealkylation of copper(II) complexes of *N*-methylporphyrins [a]

Complex	$\Delta H^{\neq}$ (kcal mol^{-1})	$\Delta S^{\neq}$ (eu)	$\Delta G^{\neq}_{298}$ (kcal mol^{-1})
Cu(N-CH$_3$TPP)$^+$	16.9 ± 1.0	-13.1 ± 2.9	20.8 ± 1.9
Cu(N-CH$_3$DP)$^+$	15.3 ± 0.3	-20.2 ± 0.9	21.3 ± 0.6
Cu(N-CH$_3$TPPS$_4$)$^{3-}$	15.6 ± 0.7	-17.6 ± 2.2	20.8 ± 1.3

[a] Ref. 204.

final spectra of PdTPP while in dimethylformamide there are isosbestics for initial spectra of the free base *N*-CH_3HTPP and PdTPP indicating a distinction between the rate determining steps in the two solvents [203]. For Ni(II), the *N*-CH_3TPP complex is formed relatively rapidly in CH_3CN [202] but the dealkylation reaction rate is comparable to the complexation rate in DMF [8].

The dealkylation rate is relatively insensitive to the nature of the porphyrin ring substituents [204] (Table 9), allowing direct comparison of results obtained for synthetic porphyrins and naturally-derived porphyrins. This fact, along with the finding that the dealkylation is highly sensitive to the nature of the *N*-substituent [205] and that Fe(II) is readily removed from its *N*-substituted porphyrin complexes [206] is of interest with regard to the formation of *N*-alkylporphyrins from cytochrome P-450 [208].

G. AGGREGATION REACTIONS

The interpretation of the kinetics of metalloporphyrins reactions depends on the knowledge of this state of aggregation. Most workers have used relatively low porphyrin concentrations (10^{-5} M or less) to avoid aggregation and consistency of Beer's Law plots to demonstrate monomeric behavior is often reported. Hambright and coworkers [209] have compiled data fora number of free base porphyrins which can be used as a guide but few kinetics studies have been reported. Krishnamurthy et al. compiled dimerization rates for several water soluble porphyrins and four, five and six coordinate complexes, finding that the rate and equilibrium constants for the four coordinate complexes are comparable to the free base porphyrins while five and six coordinate complexes exhibit much less tendency to dimerize in aqueous solution (a number of these data were originally reported by Pasternack and coworkers [210–212])[213]. Harris and Toppen used concentration jump and concentration/pH jump kinetics studies to determine the rates of dissociation of the cationic dimeric ferriporphyrin, Fe(III)TMPyP, and of the anionic dimeric Fe(III)TPPS$_4$ [214]. In both cases, the rate of monomer formation is saturated at low pH, suggesting a rate-limiting step involving monomer formation from a μ-hydroxo dimeric intermediate. Rate constants for dissociation of the Fe(III)TMPyP and Fe(III)TPPS$_4$ μ-hydroxo dimers are 0.16 s^{-1} and 2.7 s^{-1}, respectively. These results represent a different interpretation of the mechanism of ferriporphyrin monomerization from other reports [215,216].

The Fe(III) complex of deuteroporphyrin IX exhibits peroxidase-like activity. The reaction of Fe(III)DP with H_2O_2 oxidizes phenol and substituted phenols as a function of pH in a manner that strongly suggests a mixed oxidation state dimer as a peroxidatically active intermediate with an equilibrium constant of about 10^6 M^{-1} at pH 7.3 [217].

The Mn(III) complex of a tetrakis(ethylenediamine) derivative of protoporphyrin IX dimerizes in aqueous solution. When imidazole binds to form the diligated dimer, the dimer can dissociate without further addition of imidazole to form two monoligated monomers ($k_1 = 200$ s^{-1}, $k_r = 4.86 \times 10^6$ M^{-1} s^{-1}, $I = 0.1$ M, pH 6.5–7.5, 20°C). An additional molecule of imidazole can bind to form diligated and monoligated monomers ($k_f = 2.5 \times 10^3$ M^{-1} s^{-1}, $k_r = 5.4 \times 10^7$ M^{-1} s^{-1}) [218]. The formation constant for the bis(imidazole) dimer (1.1×10^4) is significantly less than the dimerization constant in the absence of imidazole (2×10^6, [219]).

H. SUMMARY

While many interesting aspects of metalloporphyrin kinetics are yet to be elucidated, especially involving photochemistry and metalloproteins, the past decade has seen the resolution of a number of important questions and development of a factual base for further study. The kinetics studies of metallation reactions demonstrate the importance of the reactivity of the precursor metal complex and indicate the rate determining formation of an SAT type intermediate. Acid-assisted demetallation requires multiple protonation and is sensitive to the nature of the porphyrin ring and solvent as well as the nature of the metal ion. Electron transfer reactions can be inner-sphere or outer-sphere with rates dependent on Franck–Condon barriers, but there is no unequivocal evidence for a dependence on spin state per se.

Photochemical reactions are influenced by internal heavy atom effects and the rapid intersystem crossing rates of paramagnetic metal atoms. Finally, the formation of metalloporphyrins by dealkylation of *N*-substituted metalloporphyrins is dependent on the metal atom and the *N*-substituent, as expected for a reaction with S_N1 character and on the solvent, and nucleophile as expected for a reaction with significant S_N2 character.

ACKNOWLEDGEMENTS

Several students have greatly contributed to our kinetics studies, especially Marilyn Bain-Ackerman, Jon Doi, and Debasish Kuila. I have benefited from discussions with several individuals and I thank the authors who provided me with preprints and reprints. I gratefully acknowledge support from ACS-PRF, the National Institutes of Health (CA 25427 and RR 08176) and the CUNY PSC-BHE grants program for support of the work.

REFERENCES

1 P. Hambright, in K.M. Smith (Ed.), Porphyrins and Metalloporphyrins, Elsevier, Amsterdam, 1975, pp. 233–278.
2 W. Schneider, Struct. Bonding (Berlin), 23 (1975) 123.
3 F.R. Longo, E.M. Brown, W.G. Rau and A.D. Adler, in D. Dolphin, (Ed.), The Porphyrins, Vol. 5, Academic Press, New York, 1978, pp. 459–481.
4 B.D. Berezin, Coordination Compounds of Porphyrins and Phthalocyanines, Wiley, New York, 1981.
5 M. Tsutsui and B.A. Taylor, in K.M. Smith (Ed.), Porphyrins and Metalloporphyrins, Elsevier, New York, 1975, pp. 279–316.
6 R. Pasternack, N. Sutin and D. Turner, J. Am. Chem. Soc., 98 (1976) 1908.
7 S. Funahashi, Y. Yamaguchi and M. Tanaka, Bull. Chem. Soc. Jpn., 57 (1984) 204.
8 M.J. Bain-Ackerman and D.K. Lavallee, Inorg. Chem., 18 (1979) 3358.
9 B. Shah, B. Shears and P. Hambright, Inorg. Chem., 10 (1971) 1818.
10 D.K. Lavallee and A.E. Gebala, Inorg. Chem., 13 (1974) 2004.
11 K.A. Freeman and F. Hibbert, J. Chem. Soc. Perkin Trans., 2 (1979) 1574.
12 A. Newberger and J.J. Scott, Proc. R. Soc. London, Ser. A, 213 (1952) 307.
13 R.M. Fuoss, J. Am. Chem. Soc., 80 (1958) 5059.
14 P. Hambright and P.B. Chock, J. Am. Chem. Soc., 96 (1974) 3123.
15 S.K. Cheung, L.F. Dixon, E.B. Fleischer, O.Y. Jeter and M. Krishnamurthy, Bioinorg. Chem., 2 (1973) 281.
16 R.F. Pasternack, G.C. Vogel, C.A. Skowronek, R.K. Harris and J.G. Miller, Inorg. Chem., 20 (1981) 3763.
17 J. Turay and P. Hambright, Inorg. Chem., 19 (1980) 562.
18 T.W. Swaddle, Inorg. Chem., 19 (1980) 3203.
19 F.K. Meyer, K.E. Newman and A.E. Merbach, J. Am. Chem. Soc., 101 (1979) 5588.
20 Y. Ducomun, K.E. Newman and A.E. Merbach, Helv. Chim. Acta, 62 (1979) 2511.
21 S. Funahashi, Y. Yamaguchi and M. Tanaka, Inorg. Chem., 23 (1984) 2249.
22 F.R. Longo, E.M. Brown, D.J. Quimby, A.D. Adler and M. Meotner, Ann. N.Y. Acad. Sci., 206 (1973) 420.
23 B.D. Berezin, O.I. Koifman, T.A. Koroleva and L.V. Ushakova, Russ. J. Phys. Chem., 53 (1975) 476 and references therein.
24 B.D. Berezin and O.I. Koifman, Russ. J. Phys. Chem., 45 (1971) 820.
25 E.I. Choi and E.B. Fleischer, Inorg. Chem., 2 (1963) 94.
26 E.B. Fleischer, E.I. Choi, P. Hambright and A. Stone, Inorg. Chem., 3 (1964) 1284.
27 D.A. Brisbin and G.O. Richards, Inorg. Chem., 11 (1972) 2849.
28 D.J. Kingham and D.A. Brisbin, Inorg. Chem. 9 (1970) 2034.

29 D.A. Brisbin and R.J. Bulahura, Can. J. Chem., 46 (1968) 3431.
30 S. Funahashi, K. Saito and M. Tanaka, Bull. Chem. Soc. Jpn., 54 (1981) 2695 and references therein.
31 J. James and P. Hambright, Inorg. Chem., 12 (1973) 474.
32 A. Adeyemo, A. Shamim, P. Hambright and R.F.X. Williams, Ind. J. Chem., 21A (1982) 763.
33 P. Worthington, P. Hambright, R.F.X. Williams, J. Reid, C. Burnham, A. Shamin, J. Turay, D.M. Bell, R. Kirkland, R.G. Little, N. Datta-Gupta and B. Eisner, J. Inorg. Biochem., 12 (1980) 281.
34 P. Worthington, P. Hambright, R.F.X. Williams, M. Feldman, K.M. Smith and K. Langry, Inorg. Nucl. Chem. Lett., 16 (1980) 441.
35 B.D. Berezin and O.I. Koifman, Russ. J. Phys. Chem., 46 (1972) 24.
36 D.A. Buckingham, C.R. Clark and W.S. Webley, J. Chem. Soc., Chem. Commun., (1981) 192.
37 J. Turay and P. Hambright, Inorg. Chim. Acta, 53 (1981) L147.
38 A. Valiotti, A. Adeyemo, R.F.X. Williams, L. Ricks, J. Norton and P. Hambright, J. Inorg. Nucl. Chem., 43 (1981) 2653.
39 V. Thanabal and V. Krishnan, Polyhedron, 2 (1983) 505.
40 E.B. Fleischer and J.H. Wang, J. Am. Chem. Soc., 82 (1960) 3498.
41 B.F. Burnham and J.J. Zuckerman, J. Am. Chem. Soc., 92 (1970) 1547.
42 D.K. Lavallee and M.J. Bain-Ackerman, Bioinorg. Chem., 9 (1978) 311.
43 H. Baker, P. Hambright and L. Wagner, J. Am. Chem. Soc., 95 (1973) 5942.
44 J.P. Maquet and T. Theophanides, Can. J. Chem., 51 (1973) 219.
45 M. Berjot, L. Bernard, J.P. Macquet and T. Theophanides, J. Raman Spectrosc., 4 (1975) 3.
46 J.P. Macquet, M.M. Millard and T. Theophanides, J. Am. Chem. Soc., 100 (1978) 4741.
47 E.B. Fleischer and F. Dixon, Bioinorg. Chem., 7 (1977) 129.
48 A.M. Abeysekara, R. Grigg, J. Trocha-Grimshaw and V. Viswanatha, J. Chem. Soc., Perkin Trans. 1 (1977) 36 and 1395.
49 A. Shamin and P. Hambright, Inorg. Chem., 22 (1983) 694.
50 D.K. Lavallee and O.P. Anderson, J. Am. Chem. Soc., 104 (1982) 4707.
51 O.P. Anderson, A.B. Kopelove and D.K. Lavallee, Inorg. Chem., 19 (1980) 2101, and references therein.
52 D.K. Lavallee and G.M. Onady, Inorg. Chem., 20 (1981) 907.
53 C. Grant, Jr. and P. Hambright, J. Am. Chem. Soc., 91 (1969) 4195.
54 A. Shamin and P. Hambright, J. Inorg. Nucl. Chem., 42 (1980) 1645.
55 M. Tabata and M. Tanaka, Anal. Lett., 13 (A6) (1980) 427.
56 M. Tabata and M. Tanaka, J. Chem. Soc., Dalton Trans., 9 (1983) 1955.
57 M. Tabata and M. Tanaka, Pure Appl. Chem., 55 (1983) 151.
58 S. Funahashi, K. Saito and M. Tanaka, Bull. Chem. Soc. Jpn., 54 (1981) 2695.
59 E.B. Fleischer, Acc. Chem. Res., 3 (1970) 105.
60 J.L. Hoard in K.M. Smith (Ed.), Porphyrins and Metalloporphyrins, Elsevier, New York, 1975, pp. 317–380.
61 G.M. Cole, D.W. Doll and S.L. Holt, J. Am. Chem. Soc., 105 (1983) 4477.
62 M.B. Lowe and J.N. Phillips, Proc. XI Int. Conf. Coord. Chem., (1968) 16.
63 R.H. Schmehl, G.L. Shaw and D.G. Whitten, Chem. Phys. Lett., 58 (1978) 549.
64 H. Baker, P. Hambright, L. Wagner and L. Ross, Inorg. Chem., 12 (1973) 2200.
65 J.W. Barnes and G.D. Dorough, J. Am. Chem. Soc., 72 (1950) 4045.
66 S. Rubin, A.W. Frenkel and M.O. Kamer, J. Phys. Chem., 46 (1942) 710.

67 S. Rubin, A.W. Frenkel, M.B. Allen and P. Nahinsky, J. Am. Chem. Soc., 64 (1942) 2297.
68 D.K. Lavallee, A.B. Kopelove and O.P. Anderson, J. Am. Chem. Soc., 100 (1978) 3025.
69 J.E. Falk, Porphyrins and Metalloporphyrins, Elsevier, Amsterdam, 1964.
70 J. Reid and P. Hambright, Inorg. Chim. Acta, 33 (1979) L135.
71 R.R. Das and K.N. Rao, Inorg. Chim. Acta, 42 (1980) 227.
72 C. Stinson and P. Hambright, J. Am. Chem. Soc., 99 (1977) 2357.
73 J. Buchler, in K.M. Smith (Ed.), Porphyrins and Metalloporphyrins, Elsevier, New York, 1975, pp. 157–232.
74 P. Hambright, Inorg. Chem., 16 (1977) 2987.
75 A. Shamim and P. Hambright, Inorg. Chem., 19 (1980) 564.
76 P. Hambright and E.B. Fleischer, Inorg. Chem., 9 (1970) 1757.
77 B. Shah and P. Hambright, J. Inorg. Nucl. Chem., 32 (1970) 3420.
78 J. Nwaeme and P. Hambright, Inorg. Chem., 23 (1984) 1990.
79 A. Adeyemo, A. Valiotti, C. Burnham and P. Hambright, Inorg. Chim. Acta Lett., 54 (1981) L63.
80 A. Valiotti, A. Adeyemo and P. Hambright, Inorg. Nucl. Chem. Lett., 17 (1981) 213.
81 P. Hambright, J. Coord. Chem., 12 (1983) 297.
82 W.L. Reynolds, J. Schufman, F. Chan and R.C. Brasted, Jr., Int. J. Chem. Kinetics, 9 (1977) 777.
83 W. Reynolds, K. Koods, B. Florine, N. Johnson and K. Thielman, Int. J. Chem. Kinetics, 22 (1980) 97.
84 P. Hambright, J. Inorg. Nucl. Chem., 39 (1977) 1102.
85 P. Hambright, Inorg. Nucl. Chem. Lett. 13 (1977) 403.
86 P. Hambright, T. Gore and M. Burton, Inorg. Chem., 15 (1976) 3214.
87 D. Davis and J. Montalve, Anal. Chem., 41 (1969) 1195.
88 D.K. Lavallee, E. Huggins and S. Lee, Inorg. Chem., 21 (1982) 1552.
89 T. Small and D.K. Lavallee, 183rd Nat. A.C.S. Meeting; Seattle, WA, 1982, No. 138-INOR and Inorg. Chim. Acta Lett., submitted.
90 D.K. Lavallee, E. Huggins and S. Lee, Inorganic Reactions in S.L. Holt (Ed.), Organized Media, ACS Symp. Ser., 177 (1982) 195.
91 R. Snellgrove and R.A. Plane, J. Am. Chem. Soc., 90 (1968) 3185.
92 B. Shah, B. Shears and P. Hambright, J. Am. Chem. Soc., 93 (1971) 776.
93 T.N. Lomova and B.D. Berezin, Russ. J. Phys. Chem., 57 (1983) 563, and references therein.
94 A.N. Drobysheva and B.D. Berezin, J. Phys. Chem., 52 (1978) 1012.
95 A.N. Drobysheva, L.P. Karmanova and B.D. Berezin, Russ. J. Phys. Chem., 51 (1977) 793.
96 L.P. Karmanova, A.N. Drobysheva and B.D. Berezin, Russ. J. Phys. Chem., 51 (1977) 371.
97 T.N. Lomova and B.D. Berezin, Russ. J. Inorg. Chem., 24 (1979) 874.
98 C.-W. Hung and J. Wei, Ind. Eng. Chem. Process. Des. Dev., 19 (1980) 250 and 257.
99 J.H. Espenson and R.J. Christensen, Inorg. Chem., 16 (1977) 2561.
100 M. Krishnamurthy, Inorg. Chem., 17 (1978) 2242.
101 A. Kumar and P. Neta, J. Phys. Chem., 85 (1981) 2830.
102 E. Fleischer, R. Chapman and M. Krishnamurthy, Inorg. Chem., 18 (1979) 2156.
103 K.M. Kadish and S. Riffard, Inorg. Chim. Acta, 74 (1983) 143.
104 E.B. Fleischer, S. Jacobs and L. Mestichelli, J. Am. Chem. Soc., 90 (1968) 2527.
105 E.B. Fleischer and M. Krishnamurthy, Ann. NY Acad. Sci., 206 (1973) 32.
106 R.F. Pasternack and G.R. Parr, Inorg. Chem., 15 (1976) 3087.

107 K.R. Ashley and S. Au-Young, Inorg. Chem., 15 (1976) 1937.
108 K.R. Ashley, M. Berggen and M. Cheng, J. Am. Chem. Soc., 97 (1975) 1422.
109 M. Birus and M. Pribanic, Croat. Chem. Acta, 49 (1977) 65.
110 R.F. Pasternack, M.A. Cobb and N. Sutin, Inorg. Chem., 14 (1975) 866.
111 R.F. Pasternack, B.S. Gilles and J.P. Stromsted, Bioinorg. Chem., 8 (1978) 33.
112 Y. Hung and D.H. Busch, J. Am. Chem. Soc., 99 (1977) 4977.
113 Y. Hung, L.Y. Marton, S.C. Jackels, A.M. Tait and D.H. Busch, J. Am. Chem. Soc., 99 (1977) 4029.
114 C. Riche, A. Chiaroni, M. Perree-Fauvet and A. Gaudemer, Acta Crystallogr., Sect. B, 34 (1978) 1868 and references therein.
115 G.N. Williams and P. Hambright, Inorg. Chem., 17 (1978) 2687.
116 E.B. Fleischer and M. Krishnamurthy, J. Am. Chem. Soc., 93 (1971) 3784.
117 M. Krishnamurthy, Inorg. Chim. Acta, 25 (1977) 215.
118 K.R. Ashley, J.G. Leipoldt and V.K. Jostic, Inorg. Chem., 19 (1980) 1609.
119 J.B. Leipoldt, S.S. Basson and D.R. Rabie, J. Inorg. Nucl. Chem., 43 (1981) 3239.
120 P. O'Brien and D.A. Sweigart, Inorg. Chem., 21 (1982) 2094.
121 B.G. Kolski and R.A. Plane, J. Am. Chem. Soc., 94 (1972) 3740.
122 J.O. Satterlee, G. LaMar and T.J. Bold, J. Am. Chem. Soc., 99 (1977) 1088.
123 R.F. Pasternack, B.S. Gillies and J.R. Stahlbusch, J. Am. Chem. Soc., 100 (1978) 2613.
124 R.F. Pasternack and J.R. Stahlbush, J. Chem. Soc. Chem. Commun., (1977) 106.
125 A.N. Thompson and M. Krishnamurthy, Inorg. Chim. Acta, 34 (1979) 45.
126 D. Burdige and D.A. Sweigart, Inorg. Chim. Acta, 28 (1978) L131.
127 C.E. Holloway, D.V. Stynes and C.P.J. Vuik, J. Chem. Soc. Dalton Trans., (1982) 95.
128 F. Pomposo, D. Carruthers and D.V. Stynes, Inorg. Chem., 21 (1982) 4245.
129 R.V. Snyder and G. LaMar, J. Am. Chem. Soc., 98 (1976) 4419.
130 J.C. Swartz, M.A. Stanford, J.N. Moy, B.M. Hoffman and J.S. Valentine, J. Am. Chem. Soc., 101 (1979) 3396.
131 M.A. Stanford, J.C. Swartz, T.E. Phillips and B.M. Hoffman, J. Am. Chem. Soc., 102 (1980) 4492.
132 E.J. Rose, P.N. Venkatasubramanian, J.C. Swartz, R.D. Jones, F. Basolo and B.M. Hoffmann, Proc. Natl. Acad. Sci., USA, 79 (1982) 5742.
133 M. Inamo, S. Funahashi and M. Tanaka, Inorg. Chim. Acta, 76 (1983) L93.
134 E.F. Caldin and J.P. Field, J. Chem. Soc., Faraday Trans., 78 (1982) 1923.
135 N. Sutin, Acc. Chem. Res., 15 (1982) 275 and references therein.
136 J.R. Winkler, D.G. Nocera, K.M. Nocera, K.M. Yocom, E. Bordignon and H.B. Gray, J. Am. Chem. Soc., 104 (1982) 5798.
137 R.S. Wade, M.L. Chu and C.E. Castro, in T.E. King, S.H. Mason and M. Morrison (Eds.), Oxidases and Related Redox Systems, Pergamon Press, New York, 1982, pp. 61–83 and references therein.
138 S. Wherland and H.B. Gray, in D. Dolphin (Ed.), Biological Aspects of Inorganic Chemistry, John Wiley, New York, 1977, pp. 290–368.
139 K.M. Kadish, M.M. Morrison, L.A. Constant, L. Dickens and D.G. Davis, J. Am. Chem. Soc., 98 (1967) 8387.
140 K.M. Kadish and G. Larson, Bioinorg. Chem., 7 (1977) 95.
141 L.A. Constant and D.G. Davis, Electroanal. Chem. Interfacial Electrochem., 74 (1976) 85.
142 S. Ni, L. Dickens, J. Tappan, L. Constant and D.G. Davis, Inorg. Chem., 17 (1978) 228.
143 K.M. Kadish and D.G. Davis, Ann. N.Y. Acad. Sci., 206 (1973) 495.
144 B.A. Feinberg, M. Gross, K.M. Kadish, R.S. Marano, S.J. Pace and J. Jordan, Bioelectrochem. Bioenerg., 2 (1975) 184.

145 K.M. Kadish and C.H. Su, J. Am. Chem. Soc., 105 (1983) 177.
146 K.M. Kadish, M. Sweetland and J.S. Chen, Inorg. Chem., 17 (1978) 2795.
147 D.F. Rohrbach, E. Deutsch, W.R. Heineman and R.F. Pasternack, Inorg. Chem., 16 (1977) 2650.
148 R.F. Pasternack, Inorg. Chem., 15 (1976) 643.
149 R.D. Chapman and E.B. Fleischer, J. Am. Chem. Soc., 104 (1982) 1757.
150 R.D. Chapman and E.B. Fleischer, J. Am. Chem. Soc., 104 (1982) 1582.
151 E.B. Fleischer and S.K. Cheung, J. Am. Chem. Soc., 98 (1976) 8381.
152 R.J. Balahura and C.P. Trivedi, Inorg. Chem., 17 (1978) 3130.
153 R.F. Pasternack and N. Sutin, Inorg. Chem., 13 (1974) 1956.
154 R.F. Pasternack, M. Cobb and N. Sutin, Inorg. Chem., 14 (1975) 866.
155 J. Reid and P. Hambright, Inorg. Chem., 17 (1978) 2329.
156 P. Worthington and P. Hambright, Inorg. Chim. Acta, 46 (1980) L87.
157 R.F. Pasternack and E.B. Spiro, J. Am. Chem. Soc., 100 (1978) 968.
158 J.H. Ong and C.E. Castro, J. Am. Chem. Soc., 99 (1977) 6740.
159 M.L. Chu, C.E. Castro and G.M. Hathaway, Biochemistr., 17 (1978) 481.
160 C.E. Castro, G.M. Hathaway and R. Havlin, J. Am. Chem. Soc., 99 (1977) 8032.
161 J.C. Oxley and D.L. Toppen, Inorg. Chem., 17 (1978) 3119.
162 F.L. Harris and D.L. Toppen, Inorg. Chem., 17 (1978) 74.
163 B.B. Wayland and J.C. Swartz, Inorg. Chim. Acta, 23 (1977) 221.
164 D. Solomon, P. Peretz and M. Farazzi, J. Phys. Chem., 86 (1982) 1842.
165 D.-H. Chin, J. Del Gaudio, G.N. LaMar and A.L. Balch, J. Am. Chem. Soc., 99 (1977) 5486.
166 D.-H. Chin, G.N. LaMar and A.L. Balch, J. Am. Chem. Soc., 102 (1980) 4344.
167 R.R. Durand, Jr., S.C. Bencosme, J.P. Collman and F.C. Anson, J. Am. Chem. Soc., 105 (1983) 2710.
168 R.F. Pasternack and B. Halliwell, J. Am. Chem. Soc., 101 (1979) 1026.
169 B. Halliwell and R.F. Pasternack, Biochem. Soc. Trans., 6 (1978) 1342.
170 R.F. Pasternack and W.R. Skowronek, Jr., J. Inorg. Biochem., 11 (1979) 261.
171 Y. Ilan, J. Rabani, I. Fridovich and R.F. Pasternack, Inorg. Nucl. Chem. Lett., 17 (1981) 93.
172 R.R. Durand, Jr. and F.C. Anson, J. Electroanal. Chem., Interfacial Electrochem., 134 (1982) 273.
173 S. Kiyotaka and F.C. Anson, J. Phys. Chem., 86 (1982) 2776.
174 A. Bettelheim and T. Kuwana, Anal. Chem., 51 (1979) 2257.
175 P.A. Forshey and T. Kuwana, Inorg. Chem., 22 (1983) 699.
176 J.A.R. Van Veen, J.F. Boar, C.J. Kroese, J.G.F. Coolegem, N. DeWit and H.A. Colijn, Ber. Bunsenges. Phys. Chem., 85 (1981) 693.
177 I. Mochida, A. Yasutake, H. Fujitsu and K. Takashita, J. Phys. Chem., 86 (1982) 3468.
178 R.O. Rocklin and R.W. Murray, J. Phys. Chem., 85 (1981) 2104.
179 P. Neta, J. Phys. Chem., 85 (1981) 3678.
180 J. Barad and P. Neta, J. Phys. Chem., 87 (1983) 1502.
181 K.M. Morehouse and P. Neta, J. Chem. Phys., 88 (1984) 1575.
182 S. Baral, P. Hambright and P. Neta, J. Phys. Chem., 88 (1984) 1595.
183 G. Cauzzo, G. Gennari, G. Jori and J.D. Spikes, Photochem. Photobiol., 25 (1977) 389.
184 M. Rougee, T. Ebbessen, F. Ghetti and R.V. Bensasson, J. Phys. Chem., 86 (1982) 4404.
185 V.H. Houlding, K. Kalyansundaram, M. Graetzel and L.R. Milgrom, J. Phys. Chem., 87 (1983) 3175.
186 M.C. Richoux and A. Harriman, J. Chem. Soc. Faraday Trans. I, 78 (1982) 1873.

187 M.P. Pileni, A.M. Braun and M. Graetzel, Photochem. Photobiol., 31 (1980) 423.
188 P.B. Brugger, M. Graetzel, T. Guarr and G. McLendon, J. Chem. Phys., 76 (1982) 944.
189 W. Potter and G. Levin, Photochem. Photobiol., 30 (1979) 225.
190 S.G. Ballard and D.C. Mauzerall, J. Chem. Phys., 72 (1980) 933.
191 A.B.P. Lever, B.S. Ramaswamy, S. Licoccia, J. Photochem., 19 (1982) 173.
192 E.I. Kapinus and V.P. Staryi, Teor. Eksp. Khim., 17 (1981) 35.
193 E.I. Kapinus, V.P. Staryi and I.I. DiLung, Khim. Vys. Energ., 16 (1982) 144.
194 D. Huppert, K.D. Straub and P.M. Rentzepis, Proc. Natl. Acad. Sci. USA, 74 (1977) 4139.
195 T. Kabayashi, D. Huppert, K.D. Straub and P.M. Rentzepis, J. Chem. Phys., 70 (1979) 1720.
196 Y. Liang and D.K. Negus, Chem. Phys. Lett., 84 (1981) 236.
197 M. Momenteau and D. Lavallette, J. Am. Chem. Soc., 100 (1978) 4322.
198 H.J. Ledon, M. Bonnet and D. Galland, J. Am. Chem. Soc., 103 (1981) 6209.
199 J.S. Olson, R.E. McKinnie, M. Mims and D.K. White, J. Am. Chem. Soc., 105 (1983) 1522.
200 B. Shears and P. Hambright, Inorg. Nucl. Chem. Lett., 6 (1970) 678.
201 D.K. Lavallee, Inorg. Chem., 15 (1976) 691.
202 D.K. Lavallee, Inorg. Chem., 16 (1977) 955.
203 J.D. Doi, C. Compito-Magliozzo and D.K. Lavallee, Inorg. Chem., 23 (1984) 79.
204 D. Kuila and D.K. Lavallee, Inorg. Chem., 22 (1983) 1095.
205 D. Kuila and D.K. Lavallee, Inorg. Chem., 23 (1984) in press.
206 D.K. Lavallee, J. Inorg. Biochem., 16 (1982) 135.
207 J.D. Doi, D.K. Lavallee, S.C. Srivastava, T. Prach, P. Richards and R.A. Fawwaz, Int. J. Appl. Rad. Isotopes, 212 (1981) 120.
208 P.R. Ortiz de Montellano, H.S. Beilan, K.L. Kunze and B.A. Mico, J. Biol. Chem., 256 (1981) 4395.
209 J. Turay, P. Hambright and N. Datta-Gupta, J. Inorg. Nucl. Chem., 40 (1978) 1687.
210 R.F. Pasternack, L. Francesconi, D. Raff and E. Spiro, Inorg. Chem., 12 (1973) 2606.
211 R.F. Pasternack, P.R. Huber, P. Boyd, G. Engasser, L. Francesconi, E. Gibbs, P. Fasella, G. Cerio Venturo and L. deC. Hinds, J. Am. Chem. Soc., 94 (1972) 4511.
212 R.R. Das, R.F. Pasternack and R.A. Plane, J. Am. Chem. Soc., 92 (1970) 3312.
213 M. Krishnamurthy, J.R. Sutter and P. Hambright, J. Chem. Soc. Chem. Commun., (1975) 13.
214 F.L. Harris and D.L. Toppen, Inorg. Chem., 17 (1978) 71.
215 R.F. Pasternack, H. Lee, P. Makk and C. Spencer, J. Inorg. Nucl. Chem., 39 (1977) 1865.
216 E.B. Fleischer, J.M. Palmer, T.S. Srivastava and A. Chatterjee, J. Am. Chem. Soc., 93 (1971) 3162.
217 R.R. Das and B.S. Prabananda, J. Inorg. Nucl. Chem., 41 (1979) 1615.
218 P. Jones, D. Mantle and I. Wilson, J. Inorg. Biochem., 17 (1982) 293.
219 R.R. Das and R.A. Plane, J. Inorg. Nucl. Chem., 37 (1975) 147.