pulsed to steady state, of any species is increased by the ratio of the lifetimes, a factor that approaches 80 in this case.

It is reasonable to associate the short-lived species with a solvent orientation that leads to the strongest solute–solvent interaction. The emission spectrum of $Cr(CN)_6^{3-}$ is unusually sensitive to environment; the ratio of the 803- and 826-nm peak heights varies considerably with the counterion in $M_3Cr(CN)_6$ crystals.[8] Furthermore, the ^2E lifetime varies from 150 msec in $Cr^{3+}:K_3Co(CN)_6$ to 4 msec for $Cr(CN)_6^{3-}$ dissolved in a rigid glass at 77°K.[9] It is, therefore, not surprising to find a range of lifetimes in rigid glass environments. Apparently only a small fraction of the $Cr(CN)_6^{3-}$ is in a strongly interacting environment. The lifetimes associated with the bulk of the ions must lie in a small range near 4 msec.

In fluid solvents, the solvent relaxation time is short compared with the $Cr(CN)_6^{3-}$ lifetime, and each $Cr(CN)_6^{3-}$ can sample the entire range of solute–solvent configuration, *i.e.*, sees an average environment. This leads to exponential decay with identical lifetimes at all excitation wavelengths, and the excitation wavelength dependence of the emission spectrum disappears.

Thus, all of our results appear to be explicable in terms of environmental heterogeneity. As further support for this interpretation, we cite the properties of $[(C_4H_9)_4N]_3Cr(CN)_6$, a salt which is more soluble than $K_3Cr(CN)_6$ in nonaqueous media. This salt exhibits the same behavior as $K_3Cr(CN)_6$ in rigid H_2O, ethylene glycol, and methyl alcohol solvent. When dissolved in ethanol at low temperature, $[(C_4H_9)_4N]_3Cr(CN)_6$ shows the same lifetime behavior (nonexponential decay, wavelength dependence of decay), but the steady-state and time-resolved spectra are different in the two solvents at all excitation wavelengths, including 366 nm. This again documents the sensitivity of $Cr(CN)_6^{3-}$ emission to environment.

It is now clear that the $^4T_{2g}$ lifetime is much shorter than 50 μsec. In fact, at room temperature, the rise-time of the $^2E \rightarrow {}^4A_2$ $Cr(CN)_6^{3-}$ emission is less than 10^{-9} sec (under 398-nm excitation).[10] Thus, the $^4T_{2g}$ lifetime in $Cr(CN)_6^{3-}$ is very short and excited-state distortions, if present, do not lead to a substantial lifetime prolongation.

The red-edge effect has been observed for polar molecules dissolved in polar glasses.[7] A comparable effect with symmetric molecules was not anticipated. Care should be exercised in analyzing spectral and lifetime data in rigid media when excitation beyond the absorption maximum is involved. Precautions are particularly important under pulsed conditions when the lifetimes of "normal" and "red-edge" species may be quite different.

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Interactions of Divalent Metal Ions with Inorganic and Nucleoside Phosphates. III. Temperature Dependence of the Magnesium(II)–Adenosine 5′-Triphosphate, –Adenosine 5′-Diphosphate, and –Cytidine 5′-Diphosphate Systems

Joseph L. Banyasz and John E. Stuehr[*][1]

Contribution from the Department of Chemistry, Case Western Reserve University, Cleveland, Ohio 44106. *Received May 9, 1973*

Abstract: Kinetic data as a function of temperature over the range 5–35° are reported for the reactions of Mg^{2+} with ATP^{4-}, ADP^{3-}, and CDP^{3-}. The results for all three systems are compatible with a mechanism involving complexation with the phosphate moiety, the rate-determining step being the expulsion of a water molecule(s) from the inner hydration sphere of Mg^{2+}. The rate constants at 25° and the activation parameters are compared with the analogous quantities for the nmr water exchange rate. Our results are completely consistent with an SN1 complexation mechanism. We conclude that the activation enthalpy is a more reliable mechanistic criterion than rate constants.

Considerable evidence has been accumulated to indicate that simple metal complexing reactions occur *via* a two-step process. The first step entails the diffusion-controlled formation of an outer sphere complex in which the reaction partners are separated by an interposed solvent layer. The second step, which is rate limiting, involves the penetration of the ligand into the inner coordination sphere of the metal ion.

From a compilation[2] of a large number of results it has been concluded that many reactions involving, *e.g.*, Ni^{2+}, Co^{2+}, and Mg^{2+}, are characterized by a ligand penetration rate constant which is a function of the metal ion but is essentially independent of the nature of

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[1] National Institutes of Health Career Development Awardee (17834).

[2] M. Eigen and R. Wilkins, *Advan. Chem. Ser.*, **No. 49**, 55 (1965).

the ligand; *i.e.*, the reactions are S_N1. Most of these studies, however, were carried out at a single temperature. The activation parameters of complexation should be more sensitive to the mechanistic details of metal ion complexation.

In an earlier paper[3] we reported kinetic data at 15° for the complexation of Mg^{2+} with the adenosine, cytosine, and inorganic phosphates. We were able to correlate the results with an S_N1 mechanism, in agreement with earlier workers.[4-7] Very few studies, however, have dealt with the complexation mechanism from the point of view of activation parameters. For these reasons we are reporting here a detailed kinetic study and temperature dependence of the interactions of Mg^{2+} with adenosine 5′-triphosphate (ATP), adenosine 5′-diphosphate (ADP), and cytidine 5′-diphosphate (CDP).

Experimental Section

Materials. Solutions of ATP, ADP, and CDP were prepared daily from the sodium salts which were purchased from Boehringer and used without further purification. Stock solutions of Mg^{2+} were prepared from reagent grade nitrate obtained from Fisher Scientific and analyzed volumetrically by EDTA. The ionic strength of all solutions was maintained at 0.1 *M* with KNO_3.

Methods. All kinetic measurements were carried out with an electrical discharge temperature jump apparatus (Messanlagen Studiengesellschaft). The time range of the instrument was 1 μsec to 1 sec with a temperature rise up to 6°. Details are given elsewhere.[8]

Since activation parameters were obtained it was essential that the temperature as a function of initial capacitor voltage be accurately known. The calibration procedure was as follows. The temperature jump cell was filled with a solution of 2×10^{-5} *M* Cresol Red in 0.1 *M* Tris–hydroxyethanolamine buffer at pH 8.1 and a thermometer was inserted. The cell temperature was then slowly varied by means of a water circulator (Lauda Ultrakryomat TK-30) and the resulting transmittance changes at 575 nm were monitored on an oscilloscope (Tektronix 549). A reproducible, linear relationship was obtained between the cell temperature and the amplitude of the deflection on the oscilloscope screen. The cell was then equilibrated at a fixed temperature. The capacitor, charged to a series of values between 20 and 50 kV, was then discharged through the cell. The amplitudes of the resulting deflections on the oscilloscope screen were then measured and the corresponding temperature rises determined from the previous calibration of cell temperature *vs.* deflection amplitude. A graph of ΔT *vs.* $(kV)^2$ yielded a straight line with slope $4.2 \times 10^{-3}{}^\circ/(kV)^2$. For example, a discharge of 35 kV yielded a temperature rise of 5.1°. For all kinetic runs the initial temperature was measured within the cell by a thermometer. The temperature rise, as computed from the capacitor voltage, was added to the initial temperature to give the final value with an estimated error of ±0.3°. Temperature jumps were made at 10-min intervals to allow the test solutions to return to the initial temperature prior to each jump. The temperature range employed was approximately 5–35°.

All pH measurements were done with a Thomas 4858-L60 electrode in conjunction with a Sargent-Welch digital pH meter. Electrode standardization with Fisher Certified or NBS buffers and pH adjustment of test solutions were performed in a glass titration cell thermostated at the calculated upper temperature of the temperature jump cell by means of a Lauda K2/R refrigerated water circulator.

The metal complexing reactions were coupled to a pH indicator (Chlorophenol Red for MgADP and MgCDP, Phenol Red for MgATP) at concentrations of $2\text{–}4 \times 10^{-5}$ *M*. Blanks consisting of metal plus indicator and ligand plus indicator were found to yield only fast proton-transfer effects. In most experiments the ratio of metal to ligand was maintained close to unity with a total concentration span of 10^{-3}–10^{-5} *M*. The pH range was about 6–7 and was kept reasonably constant for each concentration dependence. In all cases the proton-transfer reactions were much faster than metal complexation and could be treated as preequilibria.

Treatment of Data. The relaxation times were computed from at least three oscilloscope traces, photographed with a Polaroid camera. All traces were manually enlarged on graph paper and then plotted on semilog paper to determine the relaxation times. The equilibrium concentrations and appropriate concentration functions were computed by means of a Univac 1108 computer.

The interpretation of the data requires the computation of the outer sphere association constant (K_{os}), which describes the formation of the encounter complex MWL in which the metal ion and ligand are separated by a solvent molecule. The value of K_{os} depends on the charges of the reactants, the ionic strength, and specific medium effects such as the binding of supporting electrolyte to the free ligand. In the presence of KNO_3 the effective outer sphere association constant K_{os}' may be written

$$K_{os}' = K_{os}^0 \pi(\gamma)/Q \qquad (1)$$

where $K_{os}^0 = (4\pi a^3 N/3000) \exp(-Z)$, $\pi(\gamma) = \gamma_M\gamma_L/\gamma_{MWL}$, and $Q = 1 + K_{KL}\bar{K}$. The quantity *Z* is given by $z_M z_L e_0^2/\epsilon kTa$; the parameter *a* is the distance of closest approach in the outer sphere complex, *N* is Avogadro's number, and ϵ is the bulk dielectric constant. The factor *Q* takes into account the fact that K^+ competes with Mg^{2+} for the free ligand (*e.g.*, ATP^{4-}); K_{KL} is the association constant of the potassium ligand complex, and $\bar{K}$ the equilibrium K^+ concentration (≈0.1 *M*). The values of K_{os}' used in this work are 80 and 34 M^{-1} for the tri- and diphosphates, respectively. The computation of these values has been discussed in detail previously.[3]

Since this paper deals with activation parameters it is also necessary to consider the temperature dependence of K_{os}'. Differentiation of eq 1 with respect to temperature yields

$$\Delta H_{os}' = \Delta H_{os}^\circ - \frac{K_{KL}\bar{K}}{1 + K_{KL}\bar{K}} \Delta H_{KL} + RT^2 \frac{d \ln \pi(\gamma)}{dT} \qquad (2)$$

where

$$\Delta H_{os}^\circ = ZRT\left[1 + \frac{1}{\epsilon}T\left(\frac{\partial \epsilon}{\partial T}\right)_P\right]$$

The values of ΔH_{os}° computed from eq 2 are 2.3 (MgATP) and 1.8 kcal/mol (MgADP and MgCDP) for $-z_M z_L = 8$ and 6, respectively. The second and third terms on the right-hand side of eq 2 can be evaluated from experimental data.

Rechnitz[9] has found that, within experimental error, only one potassium ion binds to ATP^{4-}; as a consequence, it can be readily shown that the apparent stability constant for MgATP (K'_{MgATP}) in a KNO_3 solution is given by

$$K'_{MgATP} = K_{MgATP}/(1 + K_{KATP}\bar{K}) \qquad (3)$$

where K_{MgATP} is the true stability constant. Differentiation of eq 3 with respect to temperature yields

$$\Delta H'_{MgATP} = \Delta H_{MgATP} - \frac{K_{KATP}\bar{K}}{1 + K_{KATP}\bar{K}} \Delta H_{KATP} \qquad (4)$$

The value of K_{KATP} may be estimated to be 31 M^{-1} at $I = 0.1$ *M* and 25°, so that eq 4 becomes

$$\Delta H'_{MgATP} = \Delta H_{MgATP} - 0.76\Delta H_{KATP}$$

for $\bar{K} = 0.1$ *M*. It is thus possible to obtain an estimate of ΔH_{KATP} from a comparison of $\Delta H'_{MgATP}$ and ΔH_{MgATP}. Phillips and coworkers[10] determined the enthalpy of complexation for MgATP at several ionic strengths in tetramethylammonium chloride (TMACl). The TMA^+ ion binds to ATP^{4-} very weakly and the enthalpies determined in TMACl should, to a good approximation, be equal

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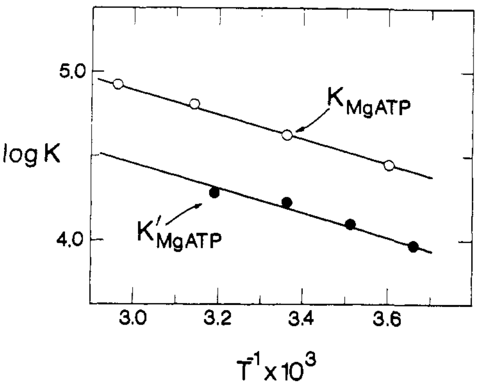

Figure 1. Temperature dependence of the binding constant for MgATP in the presence of 0.1 M TMACl (○) and 0.1 M KNO_3 (●). Data from ref 10 and 11.

to ΔH_{MgATP}. The value of $\Delta H'_{MgATP}$ in 0.1 M KNO_3 has been determined by Martell and Khan.[11] Figure 1 shows the temperature dependences of K_{MgATP} and K'_{MgATP} in 0.1 M TMACl and 0.1 M KNO_3, respectively. As expected $K'_{MgATP} < K_{MgATP}$ but, within experimental error, the slopes of the lines are identical; *i.e.*, $\Delta H'_{MgATP} \approx \Delta H_{MgATP}$ indicating that $\Delta H_{KATP} \cong 0$. This conclusion is further supported by the fact that both $\Delta H_{KP_2O_7}$ and $\Delta H_{KP_3O_{10}}$ have been found to be zero within experimental error by Wolhoff and Overbeek.[12] The second term on the right-hand side of eq 2 can therefore be neglected.

The third term on the right-hand side of eq 2 contains the temperature dependence of log $\pi(\gamma)$. The value of $\partial \log \gamma_{\pm}/\partial T$ is on the order of -2×10^{-4} deg^{-1} at $I = 0.1$ M for strong electrolytes.[13] Thus, one can calculate $RT^2 d \ln \pi(\gamma)/dT$ to be on the order of -0.08 kcal/mol, *i.e.*, negligible compared with $\Delta H_{os}°$. In support of this conclusion, the data of Phillips and coworkers[11] indicate that the value of ΔH_{MgATP} is insensitive to ionic strength. Equation 2 can, therefore, be reduced to $\Delta H_{os}' = \Delta H_{os}°$ without incurring serious error.

Results

The kinetics of the reactions of Mg^{2+} with ATP^{4-}, ADP^{3-}, and CDP^{3-} were studied as a function of temperature in the range 5–35°. Table I shows the values of the equilibrium constants used in this work as well as the corresponding enthalpies. Frey and Stuehr[14] have shown that the stability constants for MgADP and MgCDP are identical within experimental error at 15° in 0.1 M KNO_3. Further determinations indicated that this was also the case at 25°. Hence, the same stability constant and enthalpy values were used for MgADP and MgCDP. The thermodynamic parameters associated with the terminal phosphate ionization (pK_{a2}) were also found to be identical for CDP and ADP. Since pK_{a1} involves ionization of a ring nitrogen, its thermodynamic parameters are different for ADP and CDP, reflecting different ring structures. The values of pK_{a1} and ΔH_{a1} for CDP are shown in Table I.

Table I. Equilibrium Constants and Enthalpies for Metal Complex and Ligand Ionization Equilibria in 0.1 M KNO_3 at 25°

	ATP[a]	ADP[b]	CDP
pK_{a1}	4.05	3.99	4.46
ΔH_{a1}, kcal/mol	4.1	4.8	5.5
pK_{a2}	6.53	6.44	
ΔH_{a2}, kcal/mol	0.5	1.2	
log K_{MgL}	4.22	3.17	
ΔH_{MgL}, kcal/mol	2.6	3.6	
log K_{MgHL}	2.24	1.64	
ΔH_{MgHL}, kcal/mol	3.4	3.9	

[a] Reference 11; see also ref 14. [b] M. T. Kahn and A. Martell, *J. Amer. Chem. Soc.*, **89**, 5585 (1967); see also ref 14.

The relaxation times observed were in the range 0.1–1 msec. Table II gives the measured relaxation times as

Table II. Kinetic Results[a]

T, °C	M^0,[b] $M \times 10^4$	L^0,[b] $M \times 10^4$	pH[c]	f(C), $M \times 10^4$	τ^{-1},[d] sec^{-1} $\times 10^{-3}$
(1) MgATP					
9.2	19.0	22.7	7.15	7.42	3.8
9.2	9.40	11.2	7.10	4.83	2.7
9.2	4.70	5.60	7.13	3.18	2.0
9.2	2.40	2.80	7.08	2.06	1.5
$k_f = (5.0 \pm 0.6) \times 10^6$ M^{-1} sec^{-1}, $k_r = (3.8 \pm 0.4) \times 10^2$ sec^{-1}					
16.3	14.2	15.9	7.11	5.69	4.9
16.3	6.80	7.50	7.10	3.69	3.6
16.3	3.33	3.68	7.16	2.41	2.4
16.3	1.67	1.84	7.18	1.55	1.8
$k_f = (8.0 \pm 0.7) \times 10^6$ M^{-1} sec^{-1}, $k_r = (5.5 \pm 0.5) \times 10^2$ sec^{-1}					
25.0	9.80	11.2	7.15	4.41	6.7
25.0	4.90	5.60	7.14	2.93	4.4
25.0	2.38	2.72	7.11	1.87	3.3
25.0	1.19	1.36	7.12	1.19	2.9
25.0	0.60	0.68	7.11	0.67	1.9
$k_f = (1.3 \times 0.2) \times 10^7$ M^{-1} sec^{-1}, $k_r = (8 \pm 1) \times 10^2$ sec^{-1}					
35.1	4.90	4.60	7.21	2.67	9.3
35.1	2.50	2.30	7.22	1.77	6.9
$k_f = (3.0 \pm 0.3) \times 10^7$ M^{-1} sec^{-1}, $k_r = (1.55 \pm 0.15) \times 10^3$ sec^{-1}					
(2) MgADP					
6.4	30.6	4.06	5.90	29.6	6.2
6.4	14.3	4.06	5.90	14.0	3.5
6.4	6.12	4.06	5.90	6.34	2.6
6.4	14.3	6.09	5.90	13.9	3.6
6.4	30.6	10.2	5.90	28.3	5.9
6.4	12.2	10.2	5.90	11.9	3.7
$k_f = (1.55 \pm 0.2) \times 10^6$ M^{-1} sec^{-1}, $k_r = (1.6 \pm 0.2) \times 10^3$ sec^{-1}					
12.2	10.0	10.9	6.51	10.7	4.0
12.2	7.50	8.17	6.49	8.47	4.6
12.2	4.95	5.40	6.52	6.06	3.1
12.2	2.40	2.60	6.52	3.22	2.8
$k_f = (2.2 \pm 0.3) \times 10^6$ M^{-1} sec^{-1}, $k_r = (1.9 \pm 0.3) \times 10^3$ sec^{-1}					
18.4	9.80	9.50	6.51	10.0	5.6
18.4	7.17	6.95	6.52	7.86	4.4
18.4	4.63	4.48	6.48	5.45	4.5
18.4	2.31	2.24	6.50	2.98	3.3
$k_f = (3.2 \pm 0.3) \times 10^6$ M^{-1} sec^{-1}, $k_r = (2.5 \pm 0.2) \times 10^3$ sec^{-1}					
23.9	6.00	5.60	6.47	6.59	6.5
23.9	4.24	3.96	6.50	4.97	5.0
23.9	2.12	1.98	6.53	2.74	4.5
$k_f = (4.7 \pm 0.30) \times 10^6$ M^{-1} sec^{-1}, $k_r = (3.0 \pm 0.2) \times 10^3$ sec^{-1}					
30.6	7.00	7.05	6.25	7.01	12.2
30.6	3.42	3.44	5.85	3.17	8.0
$k_f = (8.9 \pm 0.9) \times 10^6$ M^{-1} sec^{-1}, $k_r = (5.4 \pm 0.6) \times 10^3$ sec^{-1}					
(3) MgCDP					
15.3	24.0	22.4	6.50	19.6	6.6
15.3	12.0	12.2	6.52	11.9	4.4
15.3	6.0	6.1	6.52	6.88	3.9
$k_f = (2.4 \pm 0.3) \times 10^6$ M^{-1} sec^{-1}, $k_r = (2.0 \pm 0.2) \times 10^3$ sec^{-1}					
24.9	41.3	41.3	6.38	26.3	19.5
24.9	20.1	20.1	6.34	16.1	13.0
24.9	4.80	4.80	6.40	5.35	6.7
24.9	2.33	2.33	6.37	2.77	5.4
$k_f = (5.8 \pm 0.6) \times 10^6$ M^{-1} sec^{-1}, $k_r = (3.9 \pm 0.4) \times 10^3$ sec^{-1}					

[a] Kinetic data obtained with 2–4 $\times 10^{-5}$ M indicator. [b] Overall concentrations. [c] a_H values converted to C_H by $\gamma_H = 0.83$. [d] ±10% deviation in the average of at least three traces.

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(11) M. T. Khan and A. Martell, *J. Amer. Chem. Soc.*, **88**, 668 (1966).

(12) J. Wolhoff and J. Overbeek, *Recl. Trav. Chim. Pays-Bas*, **78**, 759 (1959).

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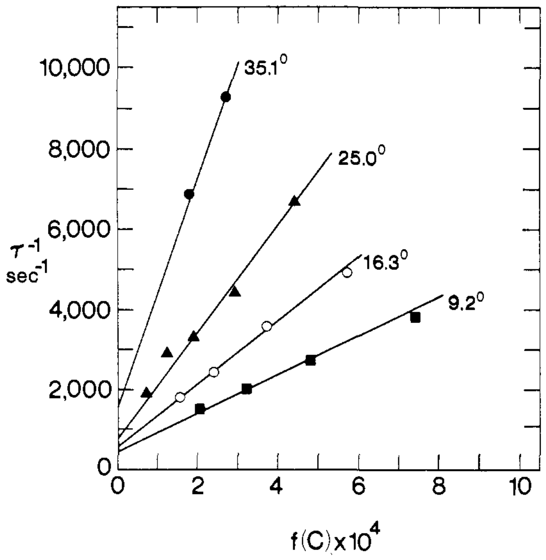

Figure 2. Dependence of τ^{-1} on f(C) (eq 5) for the MgATP system at several temperatures.

a function of concentration at the experimental temperatures as well as the concentration function f(C) which is defined later.

Mechanism. The kinetic data for MgATP, MgCDP, and MgADP are consistent with the following simple complexation mechanism

$$\begin{tikzcd}[column sep=large] \mathrm{Mg}^{2+} + \mathrm{L}^{z-} + \mathrm{H}^+ \arrow[r, shift left, "k_f"] \arrow[d, shift right, "K_a"'] & \mathrm{MgL}^{2-z} + \mathrm{H}^+ \arrow[l, shift left, "k_r"] \\ \mathrm{Mg}^{2+} + \mathrm{HL}^{1-z} \arrow[u, shift right] & \end{tikzcd} \qquad \text{(A)}$$

$$\mathrm{HIn} \xrightleftharpoons{K_{\mathrm{In}}} \mathrm{H}^- + \mathrm{In}^-$$

where L is the unprotonated ligand (ATP^{4-}, ADP^{3-}, or CDP^{3-}) and In is the unprotonated pH indicator. Eigen and coworkers,[4-6] in a study of the reactions of Mg^{2+} and Ca^{2-} with ATP and ADP, included the protonated complex MHL in the mechanism. Hammes and Levison[7] later extended the work to Ni, Co, and MnATP. They concluded that at pH ≥ 6 the species MHL could be neglected. The data presented here were taken at pH >6 and are in agreement with this conclusion. Frey, *et al.*,[3] found that at sufficiently high concentration the species $\mathrm{Mg_2L}$ must be taken into account for the reactions of Mg with ATP, ADP and CDP. The data presented here were taken at low concentrations to avoid this complication. The relaxation time for mechanism A, with the proton-transfer reactions taken as rapid preequilibria, is given by

$$\tau^{-1} = k_f\left(\frac{\overline{\mathrm{Mg}}}{1+\beta} + \bar{\mathrm{L}}\right) + k_r \qquad (5)$$

where $\beta = \bar{\mathrm{H}}/(K_a + \bar{\mathrm{L}}/(1 + \alpha))$ and $\alpha = \overline{\mathrm{In}}/(K_{\mathrm{In}} + \bar{\mathrm{H}})$. The barred quantities indicate equilibrium concentrations. The quantity $(\overline{\mathrm{Mg}}/(1 + \beta) + \bar{\mathrm{L}})$ is referred to as f(C) in the text.

A graph of τ^{-1} *vs.* f(C) yields a straight line with slope k_f and intercept k_r as shown for MgATP in Figure 2. In drawing the lines the constraint was imposed that the ratio k_f/k_r must equal the stability constant. In all cases it was possible to obtain a good fit to both the kinetic data and the stability constant. The rate constants at 25° are compared in Table III with those of other workers.

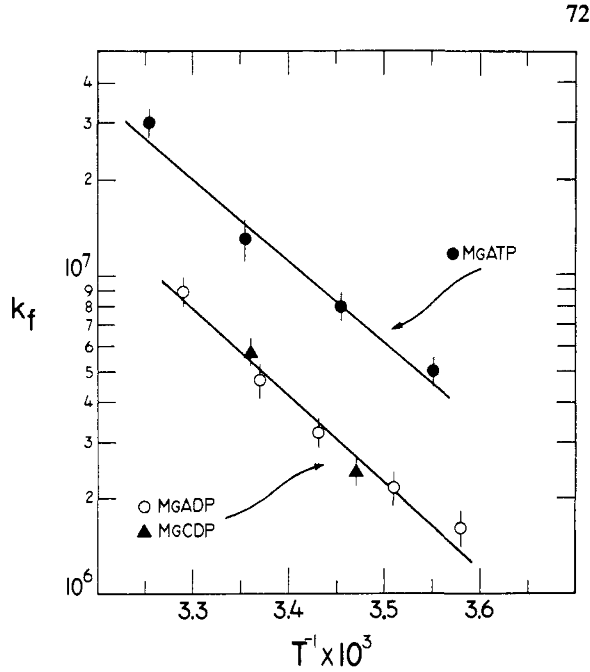

Figure 3. Variation of ln k *vs.* $1/T$ for the complexation of Mg^{2+} with three nucleotides.

Table III. Rate Constants and Activation Parameters at 25°, 0.1 M $\mathrm{KNO_3}$

	MgATP	MgADP MgCDP	Magnesium 8-hydroxy-quinoline[c]
k_f, M^{-1} sec^{-1}, present	13 ± 2	5.6 ± 0.6	
× 10^{-6}, literature	12 ± 35%[a]	3 ± 35%[b]	0.6 ± 0.05
k_r, sec^{-1}	800 ± 90	3800 ± 400	20 ±.5
K_{os},' M^{-1}	80[d]	34[d]	2.3[d]
k_{ip}, sec^{-1} × 10^{-5}	1.6	1.6	2.6[e]
$\Delta H_f^{\ddagger}$, kcal/mol	11.3 ± 1.5	12.0 ± 1.5	12.3 ± 0.6
$\Delta H_r^{\ddagger}$, kcal/mol	8.7 ± 1.5	8.5 ± 1.5	15.0 ± 2.9
$\Delta S_f^{\ddagger}$, cal/(deg mol)	12 ± 5	13 ± 5	9 ± 2
$\Delta S_r^{\ddagger}$, cal/(deg mol)	−16 ± 5	−14 ± 5	−2 ± 10

[a] Reference 4; also see ref 3. [b] Pertains to MgADP only, ref 5 and 6; see also ref 3. [c] Reference 17. [d] Computed with a = 5.3 Å. [e] Computed from data in ref 17

Temperature Dependences. The activation enthalpies were obtained from graphs of ln k/T *vs.* T^{-1} (Figure 3 shows log k values themselves for clarity), imposing the constraint that $\Delta H = \Delta H_f^{\ddagger} - \Delta H_r^{\ddagger}$. Again, in all cases excellent fits to the data were obtained within the constraint. The activation entropies were computed *via*

$$\Delta S^{\ddagger} = (\Delta H^{\ddagger} - \Delta G^{\ddagger})/T \qquad (6)$$

where $\Delta G^{\ddagger}$ is computed from the appropriate rate constant by the equation developed by Eyring. The values of the activation parameters are shown in Table III.

Discussion

The mechanism for simple metal ion complexation is usually written as

$$\mathrm{M} + \mathrm{L} \xrightleftharpoons{K_{os}} \mathrm{MWL} \xrightleftharpoons{k_{ip}} \mathrm{ML} \qquad \text{(B)}$$

where MWL and ML are the outer and inner sphere complexes, respectively. The outer sphere association constant K_{os} has been described above and k_{1p} is the ligand penetration rate constant. Ligand penetration is the rate-determining step with the formation of MWL being diffusion controlled. At sufficiently low concentrations the overall forward rate constant k_f can be written as

$$k_f = K_{os}'k_{1p} \tag{7}$$

for the present conditions.

In the case of an SN1 mechanism the rate-determining step would involve primarily the dissociation of an inner sphere water molecule with little or no bond formation between the incoming ligand and the central metal ion. In the SN2 case there would be considerable bond formation as well as bond breaking in the rate-limiting step. Thus, for an SN1 reaction, the value of k_{1p} would simply be a function of the metal ion and be relatively independent of the nature of the incoming ligand, while for the SN2 case it would be relatively strongly ligand dependent. The bulk of the kinetic evidence indicates that for most labile metal ions the value of k_{1p} is, indeed, nearly ligand independent. The observed variations could, at least in part, be due to the sometimes crude approximations of K_{os}. Basolo and Pearson[15] have also suggested several possible mechanistic explanations for the relatively small deviations that have been observed. They point out that the variations could be due to a certain amount of ligand assisted dissociation or to small changes in the water exchange rates of different outer sphere complexes. Another possibility is that various ligands have different probabilities for capturing the vacated coordination site in the five-coordinate intermediate. Nevertheless, the conclusion is that metal complexation usually proceeds by an SN1 mechanism with other effects contributing to a small extent.

As mentioned earlier, the values of K_{os}' are 80 and 34 M^{-1} at 25° and $I = 0.1\ M$ (KNO_3) for the reaction of Mg^{2+} with ATP^{4-} and the diphosphates, respectively. The experimentally determined k_f values thus yield $k_{1p} = 1.6 \times 10^5$ sec^{-1} at 25° for all three systems; this is somewhat lower than the value computed for magnesium 8-hydroxyquinoline (see Table III). Previous measurements on Mg systems by sound absorption yield $k_{1p} = 1 \times 10^5$ sec^{-1} at 20°.[2] Conversion of our values to 20° *via* the appropriate activation energies yields 1.2–1.3 × 10^5 sec^{-1}. The agreement with previous kinetic measurements is thus quite good. For water exchange, the nmr measurements of Neely and Connick yielded a k_{H_2O} value of 5.3 × 10^5 sec^{-1} for Mg^{2+} in aqueous solution at 25°.[16] The authors pointed out that differences of the same relative magnitude have been observed with other metal ions and suggest that the differences may reflect the fact that water is to some extent favored, as compared to other ligands, in the competition for a vacated coordination site. This is essentially the same explanation as the statistical argument of Basolo and Pearson.[15] The single temperature kinetic data are thus in good agreement with an SN1 mechanism.

The activation parameters for k_f can be readily obtained from eq 7. Differentiation of eq 7 with respect to temperature yields

$$\Delta H_f^{\ddagger} = \Delta H_{os}' + \Delta H_{1p}^{\ddagger} \tag{8}$$

The value $\Delta H_{1p}^{\ddagger}$ is obtained *via* eq 8 from the experimentally measured value of $\Delta H_f^{\ddagger}$ and the computed value of $\Delta H_{os}'$. The magnitude of $\Delta H_{os}'$ is somewhat dependent on the distance of closest approach parameter a but is practically constant over the range a = 5–7 Å. The parameter $\Delta S_{1p}^{\ddagger}$ is readily obtained from k_{1p} and $\Delta H_{1p}^{\ddagger}$ *via* eq 6. For an SN1 mechanism, one would expect that the value of $\Delta H_{1p}^{\ddagger}$ would be largely determined by the magnitude of $\Delta H_{H_2O}^{\ddagger}$ as measured by nmr. Small contributions from ligand assisted dissociation or small perturbations of the water exchange rate should not affect the value of $\Delta H_{1p}^{\ddagger}$ to any great extent. Steric or statistical effects would make their presence felt in the value of $\Delta S_{1p}^{\ddagger}$. Thus, for SN1 reactions one would expect that $\Delta H_{1p}^{\ddagger} \sim \Delta H_{H_2O}^{\ddagger}$ with perhaps somewhat more variability in the values of $\Delta S_{1p}^{\ddagger}$. SN2 reactions are usually considered to exhibit a reduction in the value of $\Delta H_{1p}^{\ddagger}$ and increasingly negative $\Delta S_{1p}^{\ddagger}$ values as compared with the SN1 case.

Our values of $\Delta H_{1p}^{\ddagger}$ and $\Delta S_{1p}^{\ddagger}$ are compared with those of other workers in Table IV.[16,17] The activation enthalpy for water exchange in $Mg(H_2O)_6^{2+}$ is reported to be 10.2 kcal/mol. Our own measurements yield the same number for the MgCDP and MgADP systems. The result for MgATP is somewhat lower but is within estimated error the same as the other systems (compare the slopes in Figure 3). The computed $\Delta H_{1p}^{\ddagger}$ value for magnesium 8-hydroxyquinoline is a bit high; graphing ln k *vs.* $1/T$ indicates, however, that the reported error limit[17] may be somewhat optimistic. The activation enthalpies thus support the conclusion that complexation reactions with Mg^{2+} proceed by an SN1 mechanism. There is considerable variation in the values of $\Delta S_{1p}^{\ddagger}$. The latter, however, reflect not only errors in $\Delta H^{\ddagger}$'s but variations in the rate constants from which $\Delta S^{\ddagger}$ is computed.

Table IV. Activation Parameters for the Ligand Penetration Rate Constant at 25°

	$MgATP^{2-}$	$MgADP^-$ $MgCDP^-$	Magnesium 8-hydroxyquinoline[a]	$Mg(H_2O)_6^{2+}$[c]
$\Delta H_{os}°$, kcal/mol	2.3[b]	1.8[b]	0.6[b]	
$\Delta H_{1p}^{\ddagger}$, kcal/mol	9.0 ± 1.5	10.2 ± 1.5	11.7 ± 0.6	10.2[d]
$\Delta S_{1p}^{\ddagger}$, cal/(deg mol)	−5 ± 5	0 ± 5	6 ± 2	2[e]

[a] Values computed from data in ref 17. [b] Computed for a = 5.3 Å. [c] Reference 16. [d] $\Delta H_{H_2O}^{\ddagger}$; no error limits given. [e] $\Delta S_{H_2O}^{\ddagger}$; no error limits given.

It is worth emphasizing the fact that $\Delta H_{1p}^{\ddagger}$ values can be more reliable, as indicating the type of mechanism, than k_{1p}. For example, the value of k_{H_2O} is several times larger than the ligand penetration rate constants for the systems reported in this paper. Nevertheless, the $\Delta H_{1p}^{\ddagger}$ values we calculate are indistinguishable from the reported activation enthalpy for water exchange. The differences in rate constants

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reflect differences in the entropies of activation. It is not uncommon to find systems for which "corrections" to rate constants, due to steric or statistical effects, must be made. It is often unclear how to best make these corrections. Furthermore, in order to compute k_{1p} from the measured rate constant, one must estimate K_{os}', which may be subject to substantial errors for the interaction of highly charged partners. The temperature dependence of K_{os}', on the other hand, is usually quite small. As a consequence, the numerical value of $\Delta H_{1p}^{\ddagger}$ may be a more reliable criterion of the type of substitution mechanism.

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A Comparison between Enthalpies of Formation of Antimony Pentachloride Adducts in Carbon Tetrachloride and in 1,2-Dichloroethane Solution[1]

Gerd Olofsson* and Inger Olofsson

Contribution from the Thermochemistry Laboratory, Chemical Center, University of Lund, S-220 07 Lund, Sweden. *Received April 6, 1973*

Abstract: Enthalpies of formation in carbon tetrachloride solution of molecular adducts between antimony pentachloride and eight oxygen compounds of varying donor strengths have been determined calorimetrically. They are compared with earlier reported enthalpies of adduct formation in 1,2-dichloroethane. A satisfactory linear relation was observed between the two sets of values, and none of the systems studied indicated any irregular solvation interaction in 1,2-dichloroethane.

In thermodynamic studies of donor–acceptor interactions experimental conditions are sought that will allow found values to be interpreted in terms of changes in intramolecular energies of reactant and product molecules. As gas phase reactions can be realized only in a very few cases, donor–acceptor interactions must generally be studied in solutions. It becomes a prerequisite to minimize disturbing influences of intermolecular interaction. The requirements that the solvent must keep reactants and products in solution and at the same time give negligible solute–solvent interaction contributions are conflicting, and solvents like *n*-heptane and cyclohexane, which are considered to give small solvation contributions, find a limited use in studies of systems giving polar adducts.

The good solvating properties of 1,2-dichloroethane and its chemical inertness and moderately low vapor pressure at room temperature make it a useful solvent for studies of strong donor–acceptor interactions in solution.[2–4] The usability of this solvent has, however, been questioned by Drago and coworkers[5,6] who hold the opinion that "1,2-dichloroethane should be avoided as a solvent for studies of acid–base interaction because of extensive and unpredictable contributions to the enthalpies of adduct formation from solvation." Their statement is based on discussions of formation enthalpies of donor–acceptor adducts of $SbCl_5$[5] and of hydrogen-bonded complexes of *m*-fluorophenol.[6] The discussion of solvent effects in the formation of $SbCl_5$ adducts was based on results of measurements on two systems in both CCl_4 and 1,2-$C_2H_4Cl_2$ and the remaining part on merely estimated values. Neither do the results of the study of five hydrogen-bonded systems in various solvents give evidence of "unpredictable" solvation contributions in 1,2-$C_2H_4Cl_2$. It can be seen from plots that the assumption of a linear relation between enthalpy changes measured in 1,2-$C_2H_4Cl_2$ and in CCl_4 or o-$C_6H_4Cl_2$ gives a satisfactory description of the reported results, the maximum deviation being about 1.5 kJ mol^{-1} (0.4 kcal mol^{-1}). The observed effect is in our opinion not large enough to discourage the use of 1,2-$C_2H_4Cl_2$. We found it to be of interest to compare enthalpy changes in 1,2-$C_2H_4Cl_2$ with values observed in CCl_4 for a number of adducts with as widely varying donor strength as possible to gain further information about solute–solvent interactions in 1,2-$C_2H_4Cl_2$. The use of the strong Lewis acid $SbCl_5$ as the reference acid gives complete conversion to adduct in the calorimetric experiment for a large range of strengths of interaction which reduces the experimental errors in the measured enthalpy changes.

Experimental Section

Materials. Antimony pentachloride was treated as reported.[7] It was transferred from the storage ampoules by suction in an atmosphere of dry nitrogen to 1-ml glass pipettes which were sealed. 1,2-Dichloroethane (Fisher Certified ACS grade) was fractionally distilled before use. Carbon tetrachloride (BDH, ANALAR quality) was used eithout further purification. The solvents were stored over molecular sieves 4A and the water content of the calori-

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