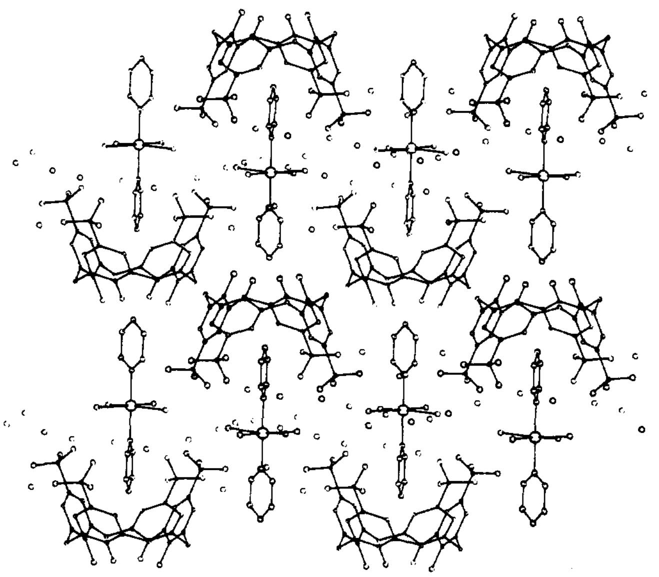

Figure 2. Structure of $[(H_2O)_4Cu(NC_5H_5)_2](H_3O)_3$[calix[4]arene sulfonate]$\cdot 9H_2O$ demonstrating the spanning of the hydrophilic layer by the copper(II) complex with one pyridine intercalated into the calixarene bilayer and the other coordinated in a second-sphere fashion by a calixarene. The Cu–N distances are 1.99 (1) and 2.00 (2) Å, and the Cu–O lengths are 2.10 (2), 2.11 (2), 2.34 (2), and 2.35 (2) Å.

is 4.3 Å, while for $C_5H_5NH^+$ in Na_4(pyridinium)[calix[4]arene sulfonate]$\cdot 8H_2O$ the value is 4.0 Å.[25] Another comparison concerns the angle made by the plane of the aromatic guest with the plane of the CH_2 carbon atoms: 54° for the pyridinium case and a more upright 74° for the nickel complex. The latter value is also reflected in the deviation of the cavity from C_{4v} symmetry for an ideal cone configuration to one of approximately C_{2v} symmetry. The S···S(trans) approaches are 8.9 and 11.4 Å, compared to 10.0 and 10.9 Å for the pyridinium complex.[25] The driving force for this is the enhanced van der Waals contacts between guest and host.

Calix[4]arene complexes of this type invariably pack into a structure that is divided into hydrophilic and hydrophobic layers. The hydrophobic layer is organized and held together by favorable van der Waals contacts and by some form of secondary bonding interaction between the sulfonate groups of the "down" calixarenes and the phenolic groups of the "up" ones. In all the previous examples this has resulted in a close-packed bilayer of calix[4]arenes.[19-21] In the title nickel complex, the second independent $[(H_2O)_5Ni(NC_5H_5)]^{2+}$ cation is seen to be intercalated into the calixarene bilayer. As is shown in Figure 1, the penetration of the intercalated transition-metal complex into the bilayer is deeper than that of the one subject to second-sphere coordination.

Very similar behavior has also been found for the $[(H_2O)_4Cu(NC_5H_5)_2]^{2+}$ cation in $[(H_2O)_4Cu(NC_5H_5)_2](H_3O)_3$[calix[4]arene sulfonate]$\cdot 9H_2O$,[26,27] as is shown in Figure 2. In this case there is only one copper-containing cation per calix[4]arene, but one pyridine is embedded in the cavity and the other is intercalated into the bilayer. The cation thus spans the hydrophilic layer.

Acknowledgment. We are grateful to the National Science Foundation for support of this work.

Supplementary Material Available: Details of X-ray crystal studies of **1** and **2** including summary of procedure and data, tables of final fractional coordinates, hydrogen atom coordinates, anisotropic thermal parameters, and bond distances and angles, and calixarene and nickel labeling schemes (18 pages); listing of observed and calculated structure factors for **1** and **2** (39 pages). Ordering information is given on any current masthead page.

---

(25) Atwood, J. L.; Orr, G. W.; Vincent, R.; Bott, S. G., unpublished results.

(26) Two equivalents of pyridine was added to 1 equiv of $Cu(NO_3)_2$ in deionized water. This solution was mixed with a second solution of 1 equiv of Na_5[calix[4]arene sulfonate] in deionized water. The resultant pale blue-green crystals of $[(H_2O)_4Cu(NC_5H_5)_2](H_3O)_3$[calix[4]arene sulfonate]$\cdot 9H_2O$ formed upon slow evaporation in a vacuum desiccator.

(27) $[(H_2O)_4Cu(NC_5H_5)_2](H_3O)_3$[calix[4]arene sulfonate]$\cdot 9H_2O$ crystallizes in the triclinic space group $P\bar{1}$ with a = 15.438 (4) Å, b = 15.727 (6) Å, c = 12.121 (9) Å, α = 112.74 (4)°, β = 102.02 (4)°, γ = 85.35 (4)°, and D_{calcd} = 1.57 g cm^{-3} for Z = 2. Refinement based on 3921 observed reflections led to a final R value of 0.125.

The Lithium Perchlorate–Diethyl Ether Rate Acceleration of the Diels–Alder Reaction: Lewis Acid Catalysis by Lithium Ion

Mark A. Forman and William P. Dailey*

Department of Chemistry, University of Pennsylvania
Philadelphia, Pennsylvania 19104-6323

Received October 1, 1990
Revised Manuscript Received January 31, 1991

The recent report by Grieco and co-workers[1] concerning the effect of 5 M lithium perchlorate–diethyl ether[2] (LPDE) on the rate of the Diels–Alder reaction was extremely stimulating. Reactions that formerly required high temperatures or high pressures for success could now be performed at room temperature and ambient pressure. We were intrigued by the very interesting suggestion that the source of this effect is due to the very high "internal pressure" of 5 M LPDE solution. As an alternative explanation we now present evidence that the rate acceleration of the Diels–Alder reaction in lithium perchlorate–diethyl ether may be due to Lewis acid catalysis, with lithium ion functioning as the Lewis acid.

If the rate enhancement is due to a "pressure" effect caused by the lithium perchlorate–diethyl ether solution, then the rates of all Diels–Alder reactions should be increased in this medium since all Diels–Alder reactions proceed with a highly negative volume of activation.[3] In contrast to the remarkable examples reported by Grieco, we have discovered that *the rate of the Diels–Alder reaction between 1,3-diphenylisobenzofuran and styrene*[4] *is unaffected by lithium perchlorate–diethyl ether solutions.*[5] A possible explanation for the increased reaction rates for some Diels–Alder reactions and not others is Lewis acid catalysis.

If lithium ion is functioning as a Lewis acid catalyst in some Diels–Alder reactions, then the second-order rate constant should depend on the lithium ion concentration. We have measured the second-order rate constants for the Diels–Alder reaction between acrylonitrile (AN) and 9,10-dimethylanthracene (DMA)[6] in the presence of various concentrations of lithium perchlorate. The reactions were carried out at 28.0 °C in ether under pseudo-first-order conditions using a large excess of dienophile and were monitored by UV spectroscopy at 398 nm by following the disappearance of DMA.[7] The second-order rate constants for the reaction are shown in Table I, and a plot of the data is shown in Figure 1. In 4.5 M LPDE, the rate enhancement is a factor of 9 over the purely thermal reaction.

---

(1) Grieco, P. A.; Nunes, J. J.; Gaul, M. D. *J. Am. Chem. Soc.* **1990**, *112*, 4595.

(2) (a) Pocker, Y.; Ciula, J. C. *J. Am. Chem. Soc.* **1989**, *111*, 4728 and references therein. (b) Braun, R.; Sauer, J. *Chem. Ber.* **1986**, *119*, 1269. (c) Herter, R.; Föhlisch, B. *Synthesis* **1982**, 976.

(3) For a review of the Diels–Alder reaction, see: Sauer, J.; Sustmann, R. *Angew. Chem., Int. Ed. Engl.* **1980**, *19*, 779.

(4) Konovalov, A. I.; Solomonov, B. N. *Zh. Org. Khim.* **1975**, *11*, 2144.

(5) At 28.0 °C the second-order rate constants in pure diethyl ether, 2.2 M $LiClO_4$–diethyl ether, and 3.6 M $LiClO_4$–diethyl ether are 1.46 × 10^{-5}, 1.41 × 10^{-5}, and 1.57 × 10^{-5} M^{-1} s^{-1}, respectively. For comparison, the second-order rate constant in pure ethanol (2.48 × 10^{-5} M^{-1} s^{-1}) shows a small increase similar to that commonly observed on going from ether to ethanol.[3]

(6) Sauer, J.; Wiest, H.; Mielert, A. *Chem. Ber.* **1964**, *97*, 3183.

(7) Typical reaction conditions were 3.0 M $LiClO_4$, 0.7 M acrylonitrile, and 2 × 10^{-4} M 9,10-dimethylanthracene.

0002-7863/91/1513-2761$02.50/0 © 1991 American Chemical Society

Table I. Second-Order Rate Constants for the Diels-Alder Reaction between 9,10-Dimethylanthracene and Acrylonitrile at 28.0 °C in Lithium Perchlorate-Diethyl Ether Solution

$k_2 \times 10^6$, $M^{-1}\,s^{-1}$	[$LiClO_4$], M	$k_2 \times 10^6$, $M^{-1}\,s^{-1}$	[$LiClO_4$], M
1.28	0	8.11	2.99
4.34	1.20	9.85	3.88
5.37	1.82	11.4	4.55

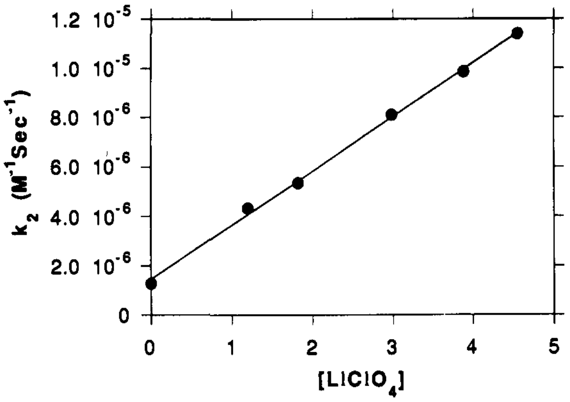

Figure 1. Second-order rate constants for the Diels-Alder reaction between acrylonitrile and 9,10-dimethylanthracene at 28.0 °C in ether versus lithium perchlorate concentration.

The linear correlation ($r = 0.999$; slope = 2.25×10^{-6}) argues that there is a first-order dependence on [$LiClO_4$]. Lewis acid catalysis of the Diels-Alder reaction is often used in organic synthesis, although relatively few detailed kinetic studies have been reported.[8] The function of the Lewis acid is well understood in terms of FMO theory, and several ab initio studies have quantified this effect.[9] Enhanced endo/exo selectivity is also observed under Lewis acid catalysis.[10] Typically, strong Lewis acids such as aluminum chloride, boron trifluoride, or tin(IV) chloride have been used, although softer species such as copper(I)[11] have also been employed, especially for sensitive substrates such as furan. In the present case we propose that lithium ion is acting as the Lewis acid[12] by complexing with the cyano group nitrogen of acrylonitrile. Gas-phase lithium ion affinities have been measured for a series of compounds, and the values for ethers, carbonyl compounds, and nitriles are all similar.[13] Ab initio calculations reproduce the experimental results.[14]

As additional examples, we have examined the reactions of fumaronitrile with DMA and isoprene with dimethyl acetylenedicarboxylate.[15] Since both fumaronitrile and dimethyl acetylenedicarboxylate have two activating groups, there are two possible sites for lithium ion complexation. For each reaction, a plot of the second-order rate constants versus [$LiClO_4$]^2 was reasonably linear. The rates for these two reactions in 4 M LPDE are accelerated by factors of 50 and 150, respectively, relative to the purely thermal reactions.

Other sources of relatively uncomplexed lithium ions behaved similarly while a perchlorate salt lacking Lewis acidity had virtually no effect on the rates. For instance, a 1.9 M solution of lithium hexafluorophosphate in ether yields a rate constant of 1.2×10^{-5} $M^{-1}\,s^{-1}$ for the Diels-Alder reaction between DMA and AN. In contrast, solutions of 1.6 M tetra-*n*-butylammonium perchlorate in tetrahydrofuran (Bu_4NClO_4 is insoluble in ether) showed no significant rate differences while comparable lithium perchlorate-tetrahydrofuran solutions showed accelerations, although much less than those observed with LPDE.[16]

In conclusion, we have presented evidence for the three systems studied here that the rate accelerations observed in LPDE are consistent with lithium ion catalysis of the Diels-Alder reaction. In addition, the accelerating effect of LPDE is not universal, suggesting that if a Diels-Alder reaction cannot be catalyzed by a Lewis acid, then LPDE may be of no value. Fortunately most reactions do not fall into this category.

Acknowledgment. M.A.F. acknowledges support by a Dissertation Dean's Fellowship from the University of Pennsylvania. W.P.D. thanks the Alfred P. Sloan Foundation for support of this research. We acknowledge frank discussions with Professors Grieco and Gajewski.

Supplementary Material Available: Listing of rate data for the reactions of fumaronitrile with DMA and isoprene with dimethyl acetylenedicarboxylate and plots of the data (2 pages). Ordering information is given on any current masthead page.

---

(8) (a) Inukai, T.; Kojima, T. *J. Org. Chem.* **1967**, *32*, 872 and references therein. (b) Kiselev, V. D.; Shakirov, I. M.; Konovalov, A. I. *Zh. Org. Khim.* **1986**, *22*, 1034. (c) Bonnesen, P. V.; Puckett, C. L.; Honeychuck, R. V.; Hersh, W. H. *J. Am. Chem. Soc.* **1989**, *111*, 6070 and references therein.

(9) (a) Birney, D. M.; Houk, K. N. *J. Am. Chem. Soc.* **1990**, *112*, 4127 and references therein. (b) Guner, O. F.; Ottenbrite, R. M.; Shillady, D. D.; Alston, P. V. *J. Org. Chem.* **1987**, *52*, 391.

(10) A report on the enhanced endo selectivity of a Diels-Alder reaction produced by concentrated lithium perchlorate-diethyl ether solutions had appeared prior to that of ref 1, although no mention was made of increased reaction rates. See ref 2b.

(11) (a) Corey, E. J.; Weinshenker, N. M.; Schaaf, T. K.; Huber, W. *J. Am. Chem. Soc.* **1969**, *91*, 5675. (b) Moore, J. A.; Partain, E. M., III. *J. Org. Chem.* **1983**, *48*, 1105. (c) Vieira, E.; Vogel, P. *Helv. Chim. Acta* **1982**, *65*, 1700.

(12) For a recent discussion on the Lewis acidity of lithium ion, see: Pushin, A. N.; Tkachenko, S. E.; Martynov, I. V. *Dokl. Akad. Nauk SSSR* **1988**, *299*, 154.

(13) (a) Staley, R. H.; Beauchamp, J. L. *J. Am. Chem. Soc.* **1975**, *97*, 5920. (b) Woodin, R. L.; Beauchamp, J. L. *J. Am. Chem. Soc.* **1978**, *100*, 501.

(14) Hehre, W. J.; Radom, L.; Schleyer, P. v. R.; Pople, J. A. *Ab Initio Molecular Orbital Theory*; Wiley: New York, 1986, p 319 f.

(15) The data is available as supplementary material.

(16) For instance, for the reaction of DMA and AN in tetrahydrofuran (THF), the second-order rate constants in pure THF, 1.6 M (n-Bu)$_4$ClO$_4$-THF, and 1.6 M LiClO$_4$-THF were 1.4×10^{-6}, 1.4×10^{-6}, and 2.2×10^{-6} $M^{-1}\,s^{-1}$, respectively.

Two-Carbon Intercalation. 4-Cyclooctenones by Tandem Application of Double Tebbe and Claisen Reactions

Christophe M. G. Philippo, Nha Huu Vo, and
Leo A. Paquette*

Evans Chemical Laboratories, The Ohio State University
Columbus, Ohio 43210

Received November 26, 1990

The intriguing structural characteristics and remarkable biological activities of natural products possessing fused eight-membered rings have made them attractive targets for synthesis.[1] However, the possibility of incorporating this uncommon carbocyclic feature into other well-known classes of pharmacologically important molecules, e.g., steroids, in an effort to potentiate biological response does not appear to have been pursued. Presumably, this is a direct consequence of the absence of a concise, reliable, and efficient scheme for effecting the enlargement of smaller functionalized rings.

---

(1) Notable examples include the following. (a) Cycloaraneosene: Kato, N.; Tanaka, S.; Takeshita, H. *Chem. Lett.* **1986**, 1989. (b) Ceroplastol II and albolic acid: Kato, N.; Takeshita, N.; Kataoka, H.; Ohbuchi, S.; Tanaka, S. *J. Chem. Soc., Perkin Trans. 1* **1989**, 165. (c) Ophiobolin C: Rowley, M.; Tsukamoto, M.; Kishi, Y. *J. Am. Chem. Soc.* **1989**, *111*, 2735. (d) Ceroplastol I: Boeckman, R. K., Jr.; Arvanitis, A.; Voss, M. E. *J. Am. Chem. Soc.* **1989**, *111*, 2737. (e) Precapnelladiene: Kinney, W. A.; Coghlan, M. J.; Paquette, L. A. *J. Am. Chem. Soc.* **1985**, *107*, 7352. (f) Poitediol: Gadwood, R. C.; Lett, R. M.; Wissinger, J. E. *J. Am. Chem. Soc.* **1984**, *106*, 3869. (g) Neolemnanyl acetate: Majetich, G.; Lowery, D.; Khetani, V. *Tetrahedron Lett.* **1990**, *31*, 51.

0002-7863/91/1513-2762$02.50/0