Ab Initio Computational Examination of Carbonyl Reductions by Borane: The Importance of Lewis Acid−Base Interactions

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Calculations of the energies and geometries of various complexes of diborane and borane with acetaldehyde, acetone, acetyl chloride, formaldehyde, methyl acetate, THF, dimethyl sulfide (DMS), and ammonia have been performed. Adducts and transition states involving methoxyborane, dimethoxyborane, and trimethyl borate have also been examined. The 6-31+G(d,p) basis set was used to optimize geometries with the Møller−Plesset (second order, MP2) perturbation method. These calculations characterize the probable intermediates and transition states involved in borane, diborane, THF·BH_3, and DMS·BH_3 reductions of carbonyl compounds as well as THF- and DMS-catalyzed diborane reductions. Four-centered transition states were located for borane reductions of acetaldehyde, acetone, acetyl chloride, formaldehyde, and methyl acetate, the accessibility of which correlates to carbonyl π-orbital energies. Diborane adducts of acetone, acetaldehyde, DMS, THF, and ammonia were also located. The ability to form such adducts depends strongly on Lewis basicity, and subsequent adduct disruption to give free borane and Lewis base−borane adducts is critical in diborane reductions. THF and DMS are predicted to disrupt completely diborane under typical conditions to give THF·BH_3 and DMS·BH_3. Borane transfer from one Lewis base to another can occur by S_N1- or S_N2-like pathways of comparable energy. Methoxyborane has a low-energy disproportionation pathway to give borane and dimethoxyborane. Dimethoxyborane is a poor reducing agent that probably does not disproportionate during reductions.

Introduction

The reaction of carbonyl compounds with diborane, borane, and borane's adducts with THF and dimethyl sulfide (DMS) have been widely exploited.[1] A recent example is the report of Corey, Bakshi, and Shibata that optically active oxazaborolidines catalyze the asymmetric reduction of ketones by H_3B·THF complex.[2] Relative to their synthetic applications, these reactions have received little mechanistic attention because of the extraordinarily reactive species involved (vide infra). This situation suggested that an ab initio computational examination might be able to illuminate details of borane and diborane reductions.[3] The results of such a study are reported revealing the importance of Lewis acid−base interactions and of the energetics of diborane disruption to give borane.

Background. In the first reported reaction of diborane with organic compounds, Brown, Schlesinger, and Burg (BSB) were surprised by the wide range of reactivity displayed by the carbonyl compounds that were the subject of their gas-phase study.[4] For example, acetaldehyde and acetone reacted within 15 min, while acetyl chloride and chloral showed no reaction. Methyl formate and ethyl acetate showed intermediate reactivity. A four-step mechanism was proposed: (1) dissociation of diborane to two borane molecules, either spontaneously or under influence of the carbonyl compound, (2) formation of a Lewis acid−base adduct between carbonyl compound and borane; (3) internal rearrangement of the preceding adduct creating new C−H and B−O bonds (carbonyl reduction and alkoxyborane formation), and (4) further reaction of the monoalkoxyborane (step 3) to give dialkoxyboranes. The differing carbonyl compound reactivities were correlated to the Lewis basicity of their carbonyl oxygens. The two esters merited further discussion because of an alternative mechanism wherein initial reduction took place by cleavage of the carbonyl carbon and alkoxy oxygen bond to give an aldehyde.

The reactivity trends of the BSB paper[4] were also observed during later work in THF and diglyme solutions.[5] In these papers, mechanistic discussion concerning carbonyl compounds of the above classes remained the same except for additional explanations and considerations concerning the low reactivity of esters. In the 1960 paper,[5b] resonance stabilization available to esters and not the other carbonyl compounds was suggested to inhibit transfer of hydride, and borane complexation to the alkoxy oxygen was suggested to compete with the carbonyl oxygen. In the 1970 paper,[5c] the stability of the tetrahedral intermediate formed after initial hydride transfer to the carbonyl carbon was evaluated with the conclusion that decomposition of the tetrahedral intermediate was fast relative to initial hydride transfer. It should be noted that THF solutions of diborane are actually THF solutions of THF·BH_3, as determined by a number of techniques.[1b]

In the first step of their carbonyl reduction mechanism, BSB proposed disruption of diborane to borane. The importance of this step is revealed in the gas-phase study of Kuhn and Doali on the reaction of acetone and

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[⊗] Abstract published in *Advance ACS Abstracts,* November 1, 1996.

[1] (a) Pelter, A.; Smith, K.; Brown, H. C. In *Borane Reagents*; Katritzky, A. R., Meth-Cohn, O., Rees, C. W., Eds.; Academic: San Diego, 1988. (b) Lane, C. F. *Chem. Rev.* **1976**, *76*, 773−799.

[2] (a) Corey, E. J.; Bakshi, R. K.; Shibata, S. *J. Am. Chem. Soc.* **1987**, *109*, 5551−5553. (b) For a review, see: Wallbaum, S.; Martens, J. *Tetrahedron: Asymmetry* **1992**, *3*, 1475−1504.

[3] The extensive computational work of Nevalainen on oxazaborolidine-catalyzed THF·H_3B reductions should be noted. For the most recent paper of a series, see: Nevalainen, V. *Tetrahedron: Asymmetry* **1994**, *5*, 903−908.

[4] Brown, H. C.; Schlesinger, H. I.; Burg, A. B. *J. Am. Chem. Soc.* **1939**, *61*, 673−680.

[5] (a) Brown, H. C.; Rao, B. C. S. *J. Org. Chem.* **1957**, *22*, 1135−1136. (b) Brown, H. C.; Subba Rao, B. C. *J. Am. Chem. Soc.* **1960**, *82*, 681−686. (c) Brown, H. C.; Heim, P.; Yoon, N. M. *J. Am. Chem. Soc.* **1970**, *92*, 1637−1646.

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diborane.[6] The acetone and diborane reaction is first order in each component, indicating that diborane does not spontaneously disrupt[7] and that acetone is possibly involved in its disruption. This is not surprising. The best experimental values for the dissociation of diborane are in the range 35−40 kcal/mol, and better ab initio computational values span the range 30−43 kcal/mol.[8] What is more important, diethyl ether and THF catalyze diborane reduction of acetone, and these reactions are first order in addend, first order in diborane, and zero order in acetone. An assisted disruption of diborane is the best explanation of these data; thus, adducts of diborane and the Lewis bases involved are implicated. The intermediacy of such adducts is supported by the observation of amine·B_2H_6 adducts,[9] and the importance of diborane disruption is echoed by several papers from Brown and co-workers concerning the chemistry of 9-BBN,[10] which exists as a dimer like borane.

In the second step of their carbonyl reduction mechanism, BSB proposed adduct formation between borane and the carbonyl compound being reduced. Such an adduct is simply a Lewis acid−base complex for which much experimental[11] and computational[12] support exists for carbonyl compounds and Lewis acids other than borane; the reactivity of borane makes such complexes very difficult to generate and examine. A single report can be found for the formation and observation of acetone·BH_3,[13] but other interpretations of the data presented are possible. Regardless, borane adducts of carbonyl compounds are reasonable intermediates.

In the third step of their carbonyl reduction mechanism, BSB proposed rearrangement of the adduct between borane and carbonyl compound to give an alkoxyborane. Since these two species are chemically plausible, they are presumably separated by at least a transition state. Possible transition state structures may be obtained by drawing analogies to alkene hydroboration. In the early 1960s, Brown and Zweifel proposed a four-centered transition state for this reaction.[14] Later Streitwieser proposed a three-centered transition state preceded by a π-complex.[15] Depending on the particular combination of computational method, basis set, and borane, evidence to support either of these transition states and an intermediate π-complex can be found.[16]

In the fourth and final step of their carbonyl reduction mechanism, BSB proposed further reaction of the monoalkoxyborane formed to give dialkoxyboranes. Just cause can be found for this. Efforts have been made to synthesize monoalkoxyboranes by treating some form of borane with either a large excess of a carbonyl compound[4,6] or an alcohol[17] without success. Monoalkoxyboranes have only been observed in a fast-flow, gas-phase, tubular reactor coupled to a mass spectrometer[18] or when formed in the presence of an amine.[19] Conversely, the accumulation of dialkoxyboranes during THF·BH_3 or diborane reductions of carbonyl compounds has been noted many times, signaling their weakness as reducing agents. Isolable dialkoxyboranes such as catecholborane fully support this view.[20]

In the following sections, ab initio calculations are described that fill some of the gaps in the experimental record and that further illuminate and confirm experiments concerning diborane reductions of carbonyl compounds.

Results and Discussion

Method. Calculations were performed on an IBM RS-6000 workstation implementing the Gaussian 92 suite of programs[21] with standard basis sets.[22] For all structures reported, stationary points (minima and transition states) were first located at the HF/6-31G(d) level and confirmed by frequency calculations. These structures then served as starting points for MP2(FC)/6-31+G(d,p) geometry optimizations. This latter level was chosen to model accurately the bridging hydrides and Lewis acid−base interactions where it is known that electron correlation and diffuse functions, respectively, are necessary to obtain accurate energies. Energy differences reported are at the MP2/6-31+G(d,p) level with scaled (0.9135[23]) zero-point energy corrections from the earlier HF/6-31G(d) frequency calculations. In order to better compare calculated values to experimental results, empirical

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[6] Kuhn, L. P.; Doali, J. O. *J. Am. Chem. Soc.* **1970**, *92*, 5475−5479.

[7] With diborane disruption rate-determining, the reaction with acetone would be first order in diborane only. With a subsequent reaction of free borane with carbonyl compound rate-determining, the reaction with acetone would be one-half order in diborane and first order in carbonyl compound.

[8] The ab initio method used here affords a diborane dissociation energy of 31.3 kcal/mol. See the following and references cited therein: (a) Page, M.; Adams, G.; Binkley, J. S.; Melius, C. F. *J. Phys. Chem.* **1987**, *91*, 2675−2678. (b) Bock, C. W.; Trachtman, M.; Murphy, C.; Muschert, B.; Mains, G. J. *J. Phys. Chem.* **1992**, *95*, 2339−2344.

[9] (a) Shore, S. G.; Hall, C. L. *J. Am. Chem. Soc.* **1966**, *88*, 5346−5347. (b) Shore, S. G.; Hall, C. L. *J. Am. Chem. Soc.* **1967**, *89*, 3947−3948.

[10] Three of the later and most relevant papers of this series follow: (a) Wang, K. K.; Brown, H. C. *J. Am. Chem. Soc.* **1982**, *104*, 7148−7155. (b) Brown, H. C.; Wang, K. K.; Chandrasekharan, J. *J. Am. Chem. Soc.* **1983**, *105*, 2340−2343. (c) Brown, H. C.; Chandrasekharan, J.; Wang, K. K. *J. Am. Chem. Soc.* **1983**, *48*, 3689−3692.

[11] For a review of Lewis acid−carbonyl interactions, see: Shambayati, S.; Schreiber, S. L. In *Comprehensive Organic Synthesis: Additions to C−X π Bonds, Part 1*; Schreiber, S. L., Ed.; Pergamon Press: Oxford, 1991; Vol. 1, pp 283−324.

[12] For example, see: (a) LePage, T. J.; Wiberg, K. B. *J. Am. Chem. Soc.* **1988**, *110*, 6642−6650. (b) Rauk, A.; Hunt, I. R.; Keay, B. A. *J. Org. Chem.* **1994**, *59*, 6808−6816.

[13] Fehlner, T. P. *Inorg. Chem.* **1972**, *11*, 252−256.

[14] Brown, H. C.; Zweifel, G. *J. Am. Chem. Soc.* **1960**, *82*, 4708−4712.

[15] Streitwieser, A., Jr.; Verbit, L.; Bittman, R. *J. Org. Chem.* **1967**, *32*, 1530−1532.

[16] For the most recent computational work in the area and a summary of past work, see: (a) Wang, X.; Li, Y.; Wu, Y.-D.; Paddon-Row, M. N.; Rondan, N. G.; Houk, K. N. *J. Org. Chem.* **1990**, *55*, 2601−2609. (b) Hommes, N. J. R. V. E.; Schleyer, P. v. R. *J. Org. Chem.* **1991**, *56*, 4074−4076.

[17] (a) Burg, A.; Schlesinger, H. I. *J. Am. Chem. Soc.* **1933**, *55*, 4020−4025. (b) Lehmann, W. J.; Onak, T. P.; Shapiro, I. *J. Chem. Phys.* **1959**, *30*, 1215−1218. (c) Lehmann, W. J.; Weiss, H. G.; Shapiro, I. *J. Chem. Phys.* **1959**, *30*, 1222−1226.

[18] Fehlner, T. P. *Inorg. Chem.* **1973**, *12*, 98−102.

[19] (a) Cai, D.; Tschaen, D.; Shi, Y.-J.; Verhoeven, T. R.; Reamer, R. A.; Douglas, A. W. *Tetrahedron Lett.* **1993**, *34*, 3243−3246. (b) Farfán, N.; Contreras, R. *J. Chem. Soc., Perkin Trans. 2* **1988**, 1787−1791.

[20] (a) Kabalka, G. W.; Baker, J. D., Jr.; Neal, G. W. *J. Org. Chem.* **1977**, *42*, 512−517. (b) Lane, C. F.; Kabalka, G. W. *Tetrahedron* **1976**, *32*, 981−990.

[21] Frisch, M. J.; Trucks, G. W.; Head-Gordon, M.; Gill, P. M. W.; Wong, M. W.; Foresman, J. B.; Johnson, B. G.; Schlegel, H. B.; Robb, M. A.; Replogle, E. S.; Gomperts, R.; Andres, J. L.; Raghavachari, K.; Binkley, J. S.; Gonzalez, C.; Martin, R. L.; Fox, D. J.; Defrees, D. J.; Baker, J.; Stewart, J. J. P.; Pople, J. A. *Gaussian 92*, Revision B; Gaussian, Inc.: Pittsburgh, PA, 1992.

[22] 6-31G(d): Hariharan, P. C.; Pople, J. A. *Theor. Chim. Acta* **1973**, *28*, 213−222. Francl, M. M.; Pietro, W. J.; Hehre, W. J.; Binkley, J. S.; Gordon, M. S.; DeFrees, D. J.; Pople, J. A. *J. Chem. Phys.* **1982**, *77*, 3654−3665. 6-31+G(d,p): Spitznagel, G. W.; Clark, T.; Chandrasekhar, J.; Schleyer, P. v. R. *J. Comput. Chem.* **1982**, *3*, 363−371. Clark, T.; Chandrasekhar, J.; Spitznagel, G. W.; Schleyer, P. v. R. *J. Comput. Chem.* **1983**, *4*, 294−301.

[23] Pople, J. A.; Scott, A. P.; Wong, M. W.; Radom, L. *Isr. J. Chem.* **1993**, *33*, 345−350. In several instances, ZPE corrections were obtained from MP2/6-31+G(d,p) calculations. These were scaled by 0.9646 (see preceding citation).

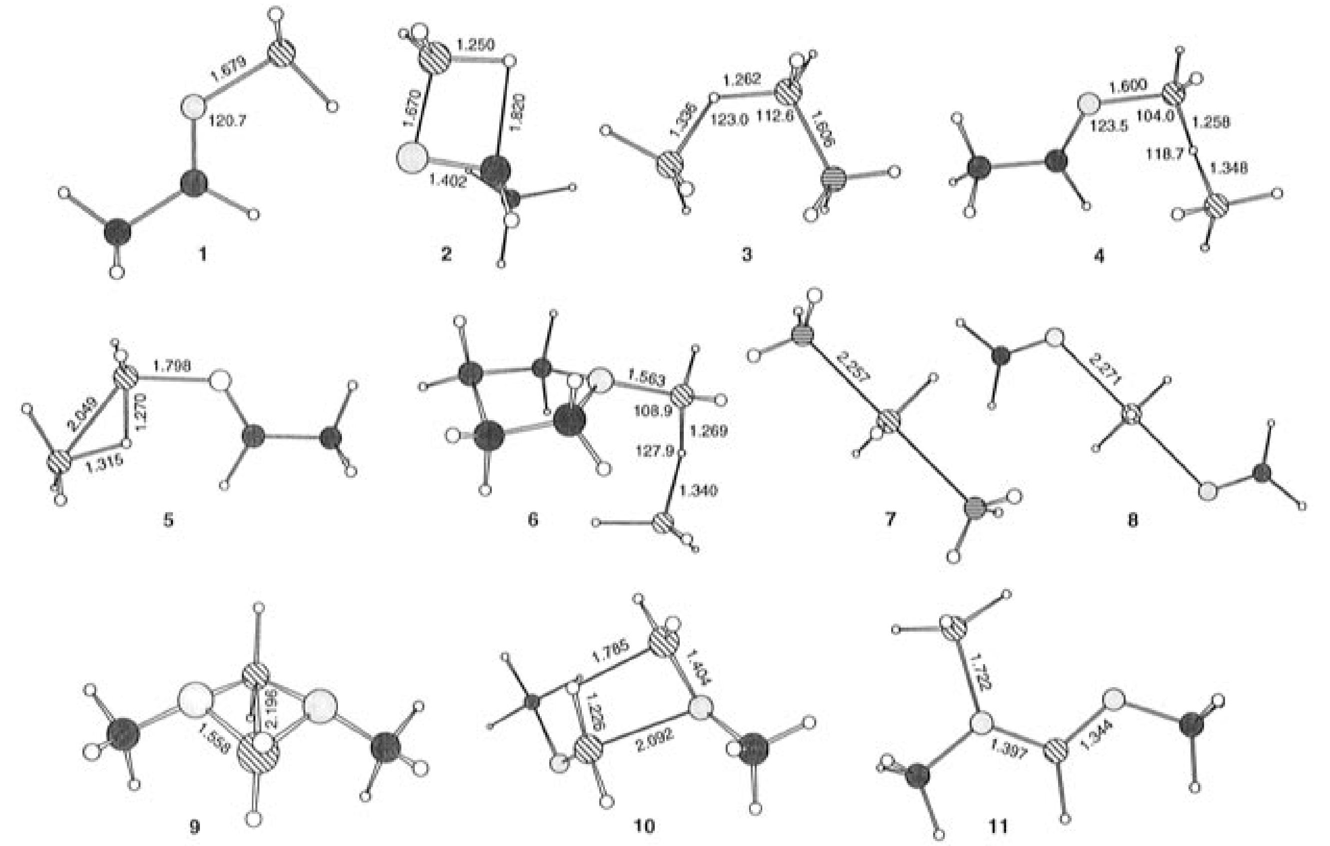


Figure 1. Minimized structures for selected compounds at the MP2/6-31+G(d,p) level. Bond lengths are in angstroms, and bond angles are in degrees.

corrections were used of (1/2)RT for each rotational or translational degree of freedom and RT whenever the work term pV arose. No corrections were made for changes in vibrational energy due to thermal effects. Normal mode analysis was performed for the imaginary frequency obtained with each transition state structure. In all cases, the behavior described in the text is consistent with these analyses. Representations of optimized structures for all numbered compounds are shown in Figure 1. The total energies and unscaled zero-point energies for all species are given in Table 1.

Borane and Carbonyl Compounds. Studies were initiated by examining the interaction of borane with five carbonyl compounds: acetaldehyde, acetone, acetyl chloride, formaldehyde, and methyl acetate. The reaction profile found for the formation of ethoxyborane from acetaldehyde and borane is representative (Figure 2). As has been previously observed computationally,[12a] combining borane with acetaldehyde is an exothermic process (−13.0 kcal/mol) to give a Lewis acid−base adduct (**1**, boron anti to methyl group) with a B−O−C bond angle of 120.7° and a B−O bond length of 1.679 Å.

A transition state for acetaldehyde hydroboration (**2**, -523 cm^{-1}) can be located. It is 15.8 and 2.8 kcal/mol higher in energy than adduct **1** and its separate components, respectively, and is presumably accessed by rearrangement of adduct **1**. The internuclear distances of **2** are consistent with an early transition state and do not depend greatly on the level of theory used. The B−O bond distance is 1.821 Å, and the hydrogen being transferred is 2.004 Å from the carbonyl carbon. Conversely, the carbonyl C−O bond and the B−H bond of the hydrogen being transferred are only slightly elongated from their equilibrium distances (0.032 and 0.034 Å, respectively). Transition state **2** is clearly four-centered. The contacts between boron and the carbon and oxygen of the carbonyl differ in length by 0.292 Å. This should be contrasted with the transition state found for the reaction of ethylene and dimethylborane, which is best described as three-centered.[16b]

Normal mode analysis indicates that ethoxyborane is provided from transition state **2**. Monoalkoxyboranes such as ethoxyborane are the expected initial products of borane reductions of aldehydes and ketones. The exothermicity of converting borane and acetaldehyde to ethoxyborane is notable (−49.4 kcal/mol) and demonstrates the tremendous reducing power of borane. This is also consistent with transition state **2** being described as early.

The interactions of acetone, acetyl chloride, formaldehyde, and methyl acetate with borane parallel that of acetaldehyde (Table 2). In all cases, Lewis acid−base complexation is exothermic, small barriers are found for hydroboration, and Lewis acid−base adducts display geometries expected from previous ab initio calculations.[12] Interestingly, there is an almost linear relationship between hydroboration transition state barrier and borane complexation energy, which are separated by 15.5 kcal/mol on average. Borane complexation energy can be viewed as a measure of electron availability for each carbonyl compound. Support for this is found in higher lying carbonyl π-electrons corresponding to increased borane complexation energies. Thus, decreasing hydroboration transition state barriers (ease of carbonyl compound reduction) correlate to increased availability of electron density. At the level of theory used here, ethylene has a π-orbital energy of 10.19 eV. This is much higher lying than the carbonyl compounds of Table 2,

Table 1. Absolute Energies (hartrees) and Zero-Point Energies (kcal/mol) for Various Species

species	sym	MP2(FC)/6-31+G(d,p)	ZPE[a]
BH_3	D_{3h}	−26.492 06	17.18[b]
CH_3CHO	$\mathrm{C_s}$	−153.389 08	37.61
$CH_3CHO{\cdot}BH_3$ (**1**)	$\mathrm{C_s}$	−179.902 60	58.88
$[CH_3CHO{\cdot}BH_3]^‡$ (**2**)	$\mathrm{C_1}$	−179.877 02	58.56
$CH_3CH_2OBH_2$	C_1	−179.965 17	61.97
CH_3COCH_3	C_{2v}	−192.583 54	56.42
$CH_3COCH_3{\cdot}BH_3$	$\mathrm{C_s}$	−219.073 94	77.71
$[CH_3COCH_3{\cdot}BH_3]^‡$	$\mathrm{C_s}$	−219.073 94	77.41
$CH_3CO_2CH_3$	$\mathrm{C_s}$	−267.637 42	60.91
$CH_3CO_2CH_3{\cdot}BH_3$	$\mathrm{C_s}$	−294.149 90	81.99
$[CH_3CO_2CH_3{\cdot}BH_3]^‡$	C_1	−294.124 55	81.83
CH_3COCl	$\mathrm{C_s}$	−612.431 46	32.06
$CH_3COCl{\cdot}BH_3$	$\mathrm{C_s}$	−638.933 12	51.10
$[CH_3COCl{\cdot}BH_3]^‡$	C_1	−638.911 50	52.74
H_2CO	C_{2v}	−114.192 63	71.14[b]
$H_2CO{\cdot}BH_3$	$\mathrm{C_s}$	−140.702 90	38.71[b]
$[H_2CO{\cdot}BH_3]^‡$	$\mathrm{C_s}$	−140.677 97	37.47[b]
NH_3	C_{3v}	−56.392 04	22.14[b]
B_2H_6	D_{2h}	−53.040 42	41.94[b]
$H_3N{\cdot}B_2H6$ (**3**)	C_1	−10.9446 54	67.61[b]
$H_3N{\cdot}BH_3$	C_{3v}	−82.928 59	45.11[b]
$CH_3CHO{\cdot}B_2H6$ (**4**)	C_1	−206.414 17	79.77
$[CH_3CHO{\cdot}B_2H_6]^‡$ (**5**)	C_1	−206.408 99	79.11
$CH_3CO_2CH_3{\cdot}B_2H_6$	C_1	−320.662 77	102.65
$[CH_3CO_2CH_3{\cdot}B_2H_6]^‡$	C_1	−320.658 41	102.35
$CH_3COCH_3{\cdot}B_2H_6$	C_1	−245.609 44	98.15
$[CH_3COCH_3{\cdot}B_2H_6]^‡$	C_1	−245.605 70	97.99
THF	C_2	−231.747 32	78.73
$THF{\cdot}B_2H_6$ (**6**)	C_1	−284.786 27	121.12
$[THF{\cdot}B_2H_6]^‡$	C_1	−284.779 37	120.58
$THF{\cdot}BH_3$	C_1	−258.271 88	100.27
$S(CH_3)_2$	C_{2v}	−477.173 61	51.17
$B_2H_6{\cdot}S(CH_3)_2$	C_1	−530.208 79	91.20
$H_3B{\cdot}S(CH_3)_2$	C_1	−503.699 88	72.01
$[H_3N{\cdot}BH_3{\cdot}NH_3]^‡$ (**7**)	D_{3h}	−139.299 57	66.51[b]
$[H_2CO{\cdot}BH_3{\cdot}H_2CO]^‡$ (**8**)	C_2	−254.891 36	55.62[b]
$(CH_3O)BH_2$	$\mathrm{C_s}$	−140.776 61	42.80
$[(CH_3O)BH_2]_2$ (**9**)	C_2	−281.574 96	89.31
$[[(CH_3O)BH_2]_2^‡$ (**10**)	C_1	−281.548 32	87.62
$(CH_3O)_2BH{\cdot}BH_3$ (**11**)	C_1	−281.567 19	87.08
$(CH_3O)_2BH$	$\mathrm{C_s}$	−255.060 07	66.67
$[(CH_3O)BH_2{\cdot}CH_3CHO]^‡$	C_1	−294.143 04	82.37
$(CH_3O)BH_2{\cdot}NH_3$	C_1	−197.187 76	69.91
$[(CH_3O)BH_2]^‡$	$\mathrm{C_s}$	−140.750 23	41.74
$[(CH_3O)_2BH{\cdot}CH_3CHO]^‡$	C_1	−408.405 27	105.43
$(CH_3O)_3B$	$\mathrm{C_s}$	−369.337 21	90.05

[a] Zero-point correction determined at HF/6-31G(d) level and reported in kcal/mol without scaling. [b] Zero-point correction determined at MP2(FC)/6-31+G(d,p) level.

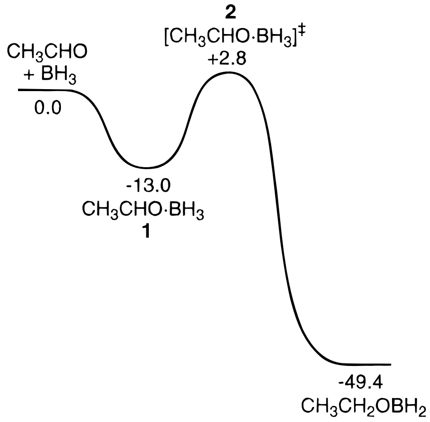


Figure 2. Profile for the reaction of borane and acetaldehyde to ethoxyborane calculated at the MP2/6-31+G(d,p) level. Relative energies have units of kcal/mol.

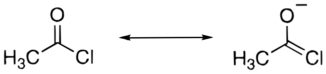
offering an explanation to transition states not being observed computationally for alkene hydroboration by borane but their being observed with carbonyl compounds.

Table 2. Carbonyl Compound π-Orbital Energies and Reaction Energetics with Borane at the MP2/6-31+G(d,p) Level

carbonyl compd	π-orbital energy (eV)	Borane complexation energy[a]	hydroboration transition state energy[a]
CH_3COCl	15.03	−7.6	+7.5
CH_2O	14.49	−9.6	+4.9
$CH_3CO_2CH_3$	14.24	−12.5	+3.3
CH_3CHO	13.63	−13.0	+2.8
CH_3COCH_3	13.17	−13.7	+1.2

[a] Energies are reported in kcal/mol, relative to isolated reactants, and ZPE corrected.

Concerning the two types of stationary points discussed, the method and basis set combinations used (HF/6-31G(d) and MP2/6-31+G(d,p)) led to only small differences in calculated geometries with one exception: the calculated B−O bond length of the Lewis acid−base adduct between borane and acetyl chloride depended strongly on the level of theory used. Hartree−Fock (HF) theory predicts a very long B−O bond using either the 6-31G(d) (2.468 Å) or 6-31+G(d,p) (2.560 Å) basis sets, typical of van der Waals bonding.[24] Møller−Plesset (MP) perturbation theory, however, predicts B−O bond lengths using the 6-31G(d) (1.747 Å) and 6-31+G(d,p) (1.740 Å) basis sets only modestly longer than those observed for the other carbonyl compound adducts with borane. This can be rationalized by MP theory better taking into account the high-energy resonance form of eq 1.[25]



(1)

Diborane and Carbonyl Compounds. To assay the effectiveness of the computational methods used here, the experimentally observed adduct between ammonia and diborane was first examined. A minimum was located (**3**) and found to be 6.9 kcal/mol more stable than its individual components.[26] Structurally, adduct **3** possesses a single boron-bridging hydrogen with markedly asymmetric B−H bonds of 1.336 and 1.262 Å (1.306 Å in diborane), suggestive of borane complexed to an electron-rich hydrogen. The borane fragment bonded directly to ammonia appears to have its Lewis acidity enhanced by having one of its hydrogens interacting with another molecule of borane. The B−N bond distance of $H_3N{\cdot}BH_3$ is 1.662 Å, and that of **3** is 1.606 Å (Δ = 0.056 Å). These data are consistent with the broadened spectra observed at temperatures above −60 °C and the B−H coupling pattern noted by Shore and Hall.[9] This good agreement between experiment and calculation indicated

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(24) See the following and references cited therein: Dvorak, M. A.; Ford, R. S.; Suenram, R. D.; Lovas, F. J.; Leopold, K. R. *J. Am. Chem. Soc.* **1992**, *114*, 108−115.

(25) This is supported by the increased C═O and decreased C−Cl bond distances on changing from HF to MP2 theory. With the 6-31G-(d) basis set, the C═O bond lengths are 1.1726 and 1.2240 Å using HF and MP2 theory, respectively. Similarly, the C−Cl bond lengths are 1.7681 and 1.734 Å for HF and MP2 theory, respectively.

(26) During the course of this work, **3** was discovered to have been computationally characterized using slightly different methods; see: (a) McKee, M. L. *J. Phys. Chem.* **1992**, *96*, 5380−5385. (b) Sakai, S. *Chem. Phys. Lett.* **1994**, *217*, 288−292. A related adduct between hydrogen sulfide and diborane has been the subject of another paper, see: Mebel, A. M.; Musaev, D. G.; Morokuma, K. *J. Phys. Chem.* **1993**, *97*, 7543−7552.

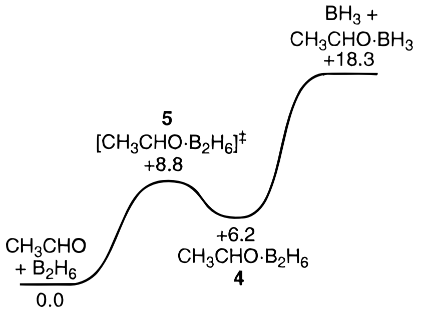


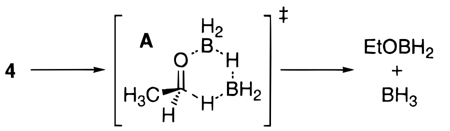
Figure 3. Profile for the reaction of diborane and acetaldehyde to the adduct of borane and acetaldehyde (**1**) and borane calculated at the MP2/6-31+G(d,p) level. Relative energies have units of kcal/mol.

that the method used here was effective and presumably applicable to other diborane adducts.

The interaction of diborane with acetaldehyde was examined next (Figure 3). As proposed by Kuhn and Doali,[6] a Lewis acid−base adduct was found between diborane and acetaldehyde (**4**). While the geometries of adducts **3** and **4** are quite similar, **4** is calculated to be 6.2 kcal/mol less stable than its constituents, which is consistent with the decreased Lewis basicity of acetaldehyde relative to ammonia. A transition state (**5**, −344 cm^{-1}) leading to **4** from diborane and acetaldehyde was also located. It resembles an S_N2 process at boron where a bridging hydrogen of diborane is the leaving group and the oxygen of acetaldehyde is the nucleophile. It is also relatively late, as the transition state B−O distance is only 0.198 Å longer than found in **4** and one of the bridging hydrogens linkages is almost completed ruptured. Dissociation of borane from **4** to give **1** requires 12.1 kcal/mol.

Considering the overall process depicted in Figure 3, an ineffective reducing agent, diborane, has been disrupted to borane, a very effective reducing agent of carbonyl compounds (vide supra), and adduct **1** with an investment of 18.3 kcal/mol. This is roughly half the energy required to disrupt diborane to two molecules of borane. The enthalpic price of disrupting diborane has been compensated for by the formation of a Lewis acid−base adduct using a path only available to a Lewis base with malleable nonbonding electrons, something not available to alkenes.

An alternative reduction mechanism proposed by Kuhn and Doali[6] also involves adduct **4** (eq 2). If **4** reorganized



(2)

to achieve six-membered transition state **A**, ethoxyborane would be afforded directly along with borane. Such a reduction mechanism would perhaps offer a more facile reduction pathway than that described already. Attempts to locate transition state **A** were unsuccessful. While this does not rule out its existence, the geometric constraints imposed by this cyclic transition state should make it relatively straightforward to find.

The interactions of acetone and methyl acetate with diborane were found to be very similar to those of acetaldehyde. Acetone and methyl acetate combine with diborane to give adducts that are 5.3 and 5.6 kcal/mol, respectively, less stable than their constituents. These two adducts are structurally similar to **3** except for the B_2H_6 unit being rotated 90° about the B−O bond axis relative to the carbonyl unit. This reorganization is presumably an adaptation to avoid steric interactions not present in acetaldehyde. Overall, the reactions of diborane with acetone and methyl acetate to give borane and their corresponding borane adducts require 17.5 and 18.8 kcal/mol, respectively. With acetaldehyde, the same process requires 18.3 kcal/mol. There are only minor differences, then, in the energy needed to generate free borane from the reaction of acetone, methyl acetate, or acetaldehyde with diborane. When combined with the knowledge that little or no additional energy is required to attain the transition states for the borane reductions of these carbonyl compounds, the experimentally slower reaction of diborane with methyl acetate is not explained.

Methyl acetate, however, differs from acetone and acetaldehyde in a fundamental way: its initial borane reduction product is a tetrahedral intermediate, $CH_3CH(OCH_3)(OBH_2)$. This tetrahedral intermediate has several reaction paths by which further reduction can take place: (1) decomposition to methoxyborane and acetaldehyde, which could then undergo further reduction; (2) internal rearrangement to effect addition of another hydrogen with methoxy group cleavage; and (3) selective reductive cleavage of the methoxy group by additional borane. If these paths are energetically more costly than tetrahedral intermediate formation, the differing carbonyl compound reactivities can be explained. This possibility was recognized by Brown, Heim, and Yoon.[5c] Their efforts to intercept tetrahedral intermediates during borane reductions of esters were unsuccessful,[27] however, leading them to conclude that tetrahedral intermediate decomposition was rapid relative to its formation. If Brown, Heim, and Yoon are correct, experimental and computational data cannot be correlated at this time.

Unlike the three carbonyl compounds already discussed, a Lewis acid−base adduct between acetyl chloride and diborane could not be located. Regardless of their relative orientation or the distance between them, acetyl chloride and diborane simply diffused apart. Given this datum and the preceding discussion, diborane reduction of acetyl chloride would be expected to be very slow or unobserved, an expectation borne out by experiment. The inability of acetyl chloride to form an adduct with diborane is presumably the result of its low Lewis basicity as indicated by the weak adduct formed between borane and acetyl chloride, weakest of all the carbonyl compounds examined (Table 2).

THF Catalysis of Diborane Reductions. As mentioned, the presence of ethers is known to catalyze diborane reductions of carbonyl compounds; therefore, the interactions of THF and DMS with diborane were examined. THF was found to form an adduct with diborane (**6**) via a transition state (−116 cm^{-1}) that resembles those already described for diborane and carbonyl compounds. Formation of adduct **6** was exothermic by 2.3 kcal/mol (Figure 4). This should be compared to the diborane adducts of acetone, acetaldehyde, and methyl acetate, whose average enthalpy of formation is endothermic by 5.7 kcal/mol. This significant enthalpic dif-

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(27) In ref 5c, ester reduction aliquots (in THF) were hydrolyzed with (2,4-dinitrophenyl)hydrazine present in the hope of forming hydrazones of hydrolytically released aldehydes. It is difficult to assess the efficacy of this method.

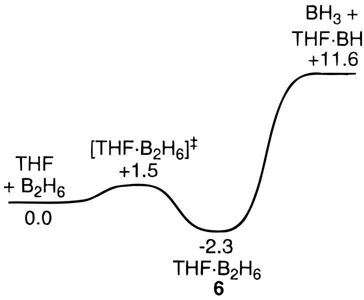

Figure 4. Profile for the reaction of diborane and THF to $H_3B{\cdot}THF$ and BH_3 calculated at the MP2/6-31+G(d,p) level. Relative energies have units of kcal/mol.

ference is not reflected in adduct geometries other than the B−O bond distance of **6** being modestly shorter (1.563 Å) than the average value for carbonyl diborane adducts (1.587 Å). Dissociation of borane from **6** to give $H_3B{\cdot}THF$ is endothermic by 13.9 kcal/mol. DMS forms a similar adduct with diborane that is 2.1 kcal/mol more stable than its constituents. Dissociation of borane from this adduct is endothermic by 12.0 kcal/mol to give $H_3B{\cdot}DMS$. Clearly, THF and DMS are stronger Lewis bases than the carbonyl compounds examined.

The acceleration of reduction rates afforded by THF can be understood by comparing the interactions of diborane with THF and acetaldehyde. The reaction of THF with diborane to give $THF{\cdot}BH_3$ and borane is endothermic by 11.6 kcal/mol with no intervening higher energy transition states or intermediates (Figure 4). Conversely, the reaction of acetaldehyde and diborane to give **1** and borane is endothermic by 18.3 kcal/mol, again with no intervening higher energy transition states or intermediates (Figure 3). Once free borane is generated, only one further barrier exists to borane reduction of carbonyl compounds like acetaldehyde and acetone, and this barrier is close to the same height as the one surmounted to generate borane, in accord with the reaction kinetics observed by Kuhn and Doali.[6] The greater Lewis basicity of THF and other ethers thus makes the disruption of diborane enthalpically more favorable, releasing borane that can then go on to reduce carbonyl compounds present, in agreement with the arguments presented by Brown.[10]

Diborane Reductions in Ethereal Solvents. The experimental behavior of diborane in coordinating solvents is mimicked computationally. The formation of 2 mol of $H_3B{\cdot}THF$ from 2 mol THF and 1 mol of diborane is predicted to be exothermic by 8.1 kcal/mol. The same reaction with DMS instead of THF is predicted to be exothermic by 11.5 kcal/mol. It should be noted that these enthalpic data overstate the favorableness of diborane dissociation by these two Lewis bases. With THF, free energy differences favor diborane and THF by 1.6 kcal/mol. This latter value is in better agreement with experiment where CH_2Cl_2 dilution of a THF solution of $THF{\cdot}BH_3$ allows diborane to be observed.[28] With DMS, free energy differences favor $DMS{\cdot}BH_3$ by 2.1 kcal/mol.

What then makes $THF{\cdot}BH_3$ and $DMS{\cdot}BH_3$ effective reducing agents? Dissociation of borane from $THF{\cdot}BH_3$ and $DMS{\cdot}BH_3$ requires 19.7 and 21.4 kcal/mol, respectively. With the generation of free borane, only small amounts of additional energy are necessary once again to reach the transition states for borane reduction of carbonyl compounds. These values for borane generation are comparable to those for acetaldehyde- and THF-catalyzed diborane disruption (18.3 and 11.6 kcal/mol, respectively) yet are significantly less than that required for uncatalyzed disruption of diborane (ca. 35 kcal/mol). In addition, reductions performed with $THF{\cdot}BH_3$ would be expected to be modestly faster that those with $DMS{\cdot}BH_3$, which is supported by experiment.

As noted earlier, the reactivity differences among carbonyl compounds translate from the gas phase to THF solutions. In THF solutions, though, borane is available at levels comparable to that afforded by the combination of acetaldehyde and diborane (above), releasing carbonyl compounds from their role of diborane disruptor. While this does not change the arguments made for the relative reactivity of methyl acetate, acetaldehyde, and acetone, the low reactivity of acetyl chloride can no longer be ascribed to its inability to disrupt diborane. However, the significantly higher energy of its borane transition state and its low Lewis basicity (Table 2), which decreases its ability to compete with THF for free borane, are probably adequate to account for its low reactivity.

The preceding discussion assumes that the coordinating solvent completely dissociates from borane before reaction can occur. An alternative view is that the coordinating solvent is still associated with borane during reactions.[29] The consensus on this matter, however, is that free borane is the reacting entity in hydroboration reactions of olefins. The transition state was located for acetaldehyde reduction with hydrated borane (HF/6-31G(d)): water is no longer associated with the borane and does not perturb the transition state geometry from that observed in its absence.

Borane Transfer. Borane transfer from one Lewis base to another can occur in two extreme manners: (1) disruption of one adduct and then combination of the resulting free borane with the other Lewis base (S_N1-like) or (2) displacement of one Lewis base from borane by another (S_N2-like). In the previous discussions, only the S_N1-like borane transfer mechanism was considered, for example, borane transfer from borane−THF to acetaldehyde.

To evaluate the likelihood of S_N2-like borane transfers, the transition states for two degenerate examples were located. In the first, the adduct of ammonia and borane is attacked by another molecule of ammonia (**7**, $-341\ cm^{-1}$). In the second, the adduct of formaldehyde and borane is attacked by another molecule of formaldehyde (**8**, $-180\ cm^{-1}$). In both cases, the expected symmetries are observed. The S_N1-like borane transfer process for $H_3N{\cdot}BH_3$ (adduct dissociation) requires 24.7 kcal/mol, while the S_N2-like process via transition state **7** requires 10.7 kcal/mol. Similarly, the S_N1-like transfer process for $H_2CO{\cdot}BH_3$ requires 9.6 kcal/mol, while the S_N2-like process via transition state **8** requires only 0.6 kcal/mol! In both these examples, an S_N2-like borane transfer process appears superior to an S_N1-like process.

While S_N2-like borane transfer between Lewis bases is facile enthalpically, entropic contributions must be considered also with such weakly bonded systems. The free energy of activation to dissociate borane from $H_2CO{\cdot}BH_3$ is expected to be slightly less than its enthalpy

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[28] Unpublished results of C. W. Lindsley and M. DiMare.

[29] For entry into this literature, see: (a) Clark, T.; Wilhelm, D.; Schleyer, P. v. R. *J. Chem. Soc., Chem. Commun.* **1983**, 606−608. (b) Brown, H. C.; Chandrasekharan, J. *Gazz. Chim. Ital.* **1987**, *117*, 517−523. (c) Tonachini, G. *Gazz. Chim. Ital.* **1988**, *118*, 149−151.

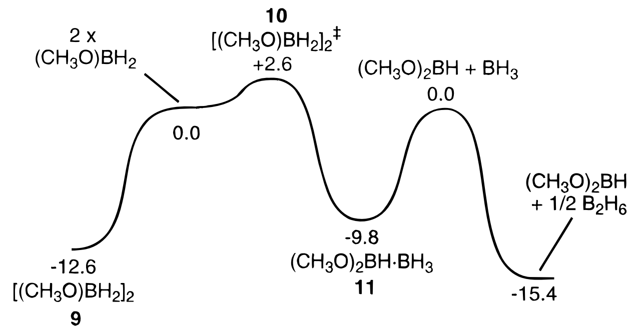

Figure 5. Profile for the disproportionation of methoxyborane to dimethoxyborane and diborane calculated at the MP2/6-31+G(d,p) level. Relative energies have units of kcal/mol.

of dissociation (9.6 kcal/mol).[30] The free energy of activation for the corresponding S_N2-like exchange is calculated to be 8.0 kcal/mol. In terms of free energy then, there is little difference between these two paths. With $H_3N{\cdot}BH_3$, the free energy of activation for the dissociative pathway is ca. 21 kcal/mol,[30] while the corresponding S_N2-like pathway is calculated to be 16.0 kcal/mol. The borane adducts of the compounds relevant to this study lie much closer to the association strength displayed by $H_2CO{\cdot}BH_3$, so S_N2-like borane transfer pathways should be considered as parallel and essentially isoenergetic means of exchanging borane between Lewis bases.[31]

Monoalkoxyboranes. Like borane, the dimerization of methoxyborane is calculated to be an exothermic process (−12.6 kcal/mol), but instead of bridging hydrogen atoms, oxygen atoms bridge to give C_2-symmetric dimer **9** (Figure 5). Since the interactions present in **9** are Lewis acid−base in nature, little or no barrier is expected in forming this dimer.[32] The B−O bonds of the dimer (1.558 Å) are considerably longer than that of the monomer (1.355 Å). Attempts to find a related dimer with one hydrogen atom and one oxygen atom bridging[33] afforded transition state **10** ($-418\ cm^{-1}$), which leads to the adduct $(CH_3O)_2BH{\cdot}BH_3$ (**11**) in a slightly endothermic process overall (+2.8 kcal/mol) from **9**. However, disruption of this adduct to dimethoxyborane and borane is an energetically facile process, leaving free borane to dimerize (as shown), for example, in an exothermic manner from both **9** and **11** (−5.6 and −2.8 kcal/mol, respectively). Monoalkoxyboranes thus appear to be prone to disproportionation.

While less facile than disproportionation, methoxyborane can behave as a reducing agent. For the methoxyborane reduction of acetaldehyde, a transition state ($-694\ cm^{-1}$) very similar to **2** was found. It is 13.6 kcal/mol higher in energy than separated reactants. Efforts to locate a Lewis acid−base adduct between these two reactants were unsuccessful. With these data, borane and methoxyborane can be compared as reducing agents. The acetaldehyde-assisted disruption of diborane (+18.3 kcal/mol) and subsequent access to the transition state for borane reduction of acetaldehyde (+2.8 kcal/mol) is endothermic by 21.2 kcal/mol overall. Thermal disruption of methoxyborane dimer (+12.6 kcal/mol) and subsequent access to the transition state for methoxyborane reduction of acetaldehyde (+13.6 kcal/mol) is endothermic by 26.2 kcal/mol overall. In the presence of THF, these energetic differences are only expected to increase because of easier borane generation (vide supra). Thus, methoxyborane is expected to be less effective than diborane as a reducing agent.

The attenuated Lewis acidity of methoxyborane relative to borane is evident. The adducts $H_3N{\cdot}BH_3$ and $H_3N{\cdot}BH_2(OCH_3)$ have heats of formation from their constituents of −24.7 and −11.1 kcal/mol, respectively, and their B−N bond lengths are 1.662 and 1.702 Å, respectively. As expected, electron donation from oxygen to boron is the principal cause of these differences and is reflected in the structure of methoxyborane. The methoxy methyl group lies in the plane defined by boron and its two hydrogens, and the B−O bond of methoxyborane is 0.012 and 0.020 Å shorter than the B−O bond distances of dimethoxyborane (averaged) and trimethyl borate, respectively. Finally, 90° rotation about the B−O bond of methoxyborane gives a transition state structure ($-464\ cm^{-1}$) that is 15.6 kcal/mol less stable and has a 0.008 Å longer B−O bond than the ground state rotamer.

The behavior described above agrees well with the inability of researchers to isolate monoalkoxyboranes. Of the two pathways discussed that lead to consumption of monoalkoxyboranes, the disproportionation pathway presumably predominates because of its lower barriers. That monoalkoxyboranes can be observed when complexed to amines is then the result of amine coordination blocking disproportionation.[34]

Dialkoxyboranes. A transition state similar to that found for the reduction of acetaldehyde by borane and methoxyborane was also found for dimethoxyborane ($-817\ cm^{-1}$). The transferring hydrogen to carbonyl carbon distance is 1.623 Å (cf. 1.755 and 2.004 Å for methoxyborane and borane, respectively), indicating a significantly later transition state.[35] This transition state is, however, 26.8 kcal/mol less stable than its components, which should be compared to the +21.2 kcal/mol overall for self-assisted acetaldehyde reduction by diborane (see above). This is consistent with dialkoxyboranes accumulating in diborane or $THF{\cdot}BH_3$ reductions. It should be noted that dialkoxyboranes are still thermodynamically potent reducing agents. Formaldehyde reduction by dimethoxyborane to trimethyl borate (−50.6 kcal/mol) is almost as exothermic as formaldehyde reduction by borane to methoxyborane (−54.1 kcal/mol).

Given the poor reducing abilities of dialkoxyboranes, their disproportionation to borane and trimethyl borate may be important. The equilibrium of eq 3 has been examined in the gas phase[17a,36] and solution.[37] Its

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[30] The difference in free energy between $H_2CO{\cdot}BH_3$ and its constituents is 0.0 kcal/mol using calculated values of $S°$ (principally due to the increased number of particles). The free energy of activation for this process is predicted to be slightly less than the enthalpic value reported. This is because the full entropic counterbalancing of enthalpic effects will not occur until very late in the reaction when ammonia and borane have dissociated (the majority of $S°$ is derived from rotational and translational terms).

[31] Whether a similar advantage exists with diborane adducts of Lewis bases is unclear. Attempts to locate an S_N2-like transition state for transferring BH_3 from $THF{\cdot}B_2H_6$ (**6**) to another Lewis base were unsuccessful.

[32] No effort was made to locate a transition state leading to dimer **9**.

[33] A hydrogen-bridged dimer of methoxyborane could not be found. No effort was made to locate it as a transition state for degenerate hydrogen exchange.

[34] The presence of THF or DMS would not be expected to modify the behavior of monoalkoxyboranes, because their coordination to a monoalkoxyborane would be weak relative to amines.

[35] Also indicating later transition states, the transferring hydrogen to boron bond lengthens in the transition states for acetaldehyde reduction: borane, 1.218 Å; methoxyborane, 1.240 Å; dimethoxyborane, 1.270 Å.

[36] Uchida, H. S.; Kreider, H. B.; Murchison, A.; Masi, J. F. *J. Phys. Chem.* **1959**, *63*, 1414−16.

$$3\,HB(OCH_3)_2 \rightleftharpoons 1/2\,B_2H_6 + 2\,B(OCH_3)_3 \qquad (3)$$

kinetics are highly variable, indicating that catalysis is critical to its achievement,[38] but generally slow (hours to days). Computationally, the disproportionation products of eq 3 are favored by 8.3 kcal/mol. That catalysis plays a role is not surprising. Dialkoxyboranes are very weak Lewis acids, binding ammonia and trimethylamine only at low temperatures (cf. borane and methoxyborane binding of ammonia above).[39,40] Therefore, the dimeric, bridged transition states necessary for ligand exchange are difficult to achieve with dialkoxyboranes. While the right side of eq 3 is favored, this equilibrium seems to have little relevance to reductions, although uncertainties remain.

The potential role of dialkoxyboranes as catalysts[41] for borane reductions merits examination since they accumulate during reductions. Much like THF, they could serve as a Lewis base to assist in the breakup of diborane. This possibility can be ruled as unlikely by comparing the complexation energies of borane to THF, acetaldehyde, and dimethoxyborane, which are −19.7, −13.0, and −9.7 kcal/mol, respectively. Thus, dimethoxyborane must be considered a less effective catalyst than the carbonyl compounds whose reductions it supposedly catalyzes since borane complexation energies correlate to diborane complexation energies. Apparently, the two oxygens of dimethoxyborane transfer enough electron density to boron to attenuate their Lewis basicity significantly.

Conclusions

Extensive ab initio calculations show that Lewis acid−base chemistry is at the heart of carbonyl reductions by diborane, borane, and borane's adducts with DMS and THF. Diborane is not an effective reducing agent until dissociated into some form of borane. The Lewis basicity of carbonyl compounds and coordinating solvents used in reductions is tightly linked to rates because of their forming borane adducts that compensate for the energy needed to dissociate diborane. Strong Lewis bases readily disrupt diborane and lead to fast reductions, while weak Lewis bases allow slow reductions at best. The use of a strong, inert Lewis base allows catalysis. Reduction occurs via a four-centered transition state. Monoalkoxyboranes are revealed as unstable species that readily disproportionate to dialkoxyboranes and diborane via a dimeric transition state. Dialkoxyboranes are shown to be poor reducing agents. The complexity and low barriers seen in diborane chemistry with carbonyl compounds should serve as a caveat to those who use them synthetically.

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(37) Pasto, D. J.; Balasubramaniyan, B.; Wojtkowski, P. W. *Inorg. Chem.* **1969**, *8*, 594−598.

(38) (a) Jeffers, P. M.; Bauer, S. H. *Inorg. Chem.* **1982**, *21*, 2516−2517. (b) Masuda and co-workers reported that 1 mol % $LiEt_3BH$ greatly accelerates the reaction of dialkoxyboranes with alcohols, see: Masuda, Y.; Nunokawa, Y.; Hoshi, M.; Arase, A. *Chem. Lett.* **1992**, 349−352.

(39) (a) McAchran, G. E.; Shore, S. G. *Inorg. Chem.* **1966**, *5*, 2044−2046. (b) Woods, W. G.; Strong, P. L. *J. Am. Chem. Soc.* **1966**, *88*, 4667−4671. (c) Young, D. E.; McAchran, G. E.; Shore, S. G. *J. Am. Chem. Soc.* **1966**, *88*, 4390−4396.

(40) Despite several attempts, an adduct of ammonia and dimethoxyborane could not be located at the HF/6-31G(d) level.

(41) Dialkoxyboranes would have to be mild catalysts not to generate a marked autocatalytic effect or be detected by earlier vapor-phase kinetics (see ref 6).