Table I

Compd	Color	Nmr[a,b] (toluene solution)
Ia	Yellow-brown	+10.0 (15 H), +202 (3 H)
Ib	Dark red-brown	+10.3 (15 H), +18.7 (3 H), +27.6 (2 H), +33.6 (2 H), +195 (2 H)
Ic	Dark brown	+10.0 (15 H), +38.2 (1 H), +126 (4 H) (+25°) +13.1 (15 H), +41.0 (1 H), +57.5 (1 H), +60.9 (1 H), +334 (2 H) (−90°)
Id	Dark red-brown	+11.6 (15 H), +22.1 (3 H), +186 (2 H)
Ie	Dark brown	+10.9 (15 H), +88.7 (2 F), 99.1 (2 F), +115.5 (1 F)
If	Dark red-brown	+10.9 (15 H), +19.3 (6 H), +190 (1 H)

[a] Pmr data in parts per million to high field of internal benzene. ^{19}F data in parts per million to high field of internal $C_6H_5CF_3$. [b] All data at +25° except where indicated.

data (Table I), which include some of the largest chemical shifts yet observed for uranium(IV) organometallics,[2] are in good accord with the above formulation. The allyl molecule is fluxional with room temperature and above, nmr spectra approaching an A_4X pattern, while the lower temperature spectra reveal collapse of the high-field peak and eventual freezing out of the *monohapto*allyl[9,10] A_2BCD pattern at −85°. Infrared spectra clearly indicate that the cyclopentadienyl rings in all molecules are π bonded.[11]

The nature of the uranium–carbon σ bond is, of course, the feature of greatest interest. In this connection, we find that these molecules all possess considerable thermal stability. In the solid state under nitrogen, there is no noticeable decomposition after days at room temperature. In toluene solution (0.054 *M*), all molecules decompose at nearly the same rate, with half-lives of 20–96 hr at 72°; at 100°, 15 hr are required to completely destroy Ib. The narrow spectrum of thermal stability we observe deviates considerably from most transition metal systems, where fluorine substitution[12,13] or prevention of β-hydrogen abstraction[13,14] greatly increases resistance to thermolysis, and this apparently anomalous behavior has prompted some mechanistic investigation. The major organic decomposition product of Ib is *n*-butane (92 ± 3% yield by glpc), with trace amounts of 1-butene (<2%) and *n*-octane (<1%) also detected. This result (especially the near absence of 1-butene) can be contrasted with a thermally less stable, well-studied *n*-butylcopper(I) system[15] where *ca.* 1:1 butane–butene was produced, suggesting that β-hydrogen elimination was operative, *i.e.*

$$M\text{-}CH_2CH_2R \longrightarrow M\text{-}H + CH_2{=}CHR$$

$$M\text{-}H + M\text{-}CH_2CH_2R \longrightarrow 2M + CH_3CH_2R$$

The uranium(IV) organometallics, however, appear to resist β elimination, and this apparently enhances the thermal stability. A homolytic, free-radical bond scission is the most reasonable alternative pathway.[13,16,17]

The polarity of the uranium–carbon bond was also investigated. All compounds react instantly with methanol, producing as the major product the known compound, $(C_5H_5)_3UOCH_3$.[19] No reaction was observed when it was attempted to bring about nucleophilic addition of $(C_5H_5)_3UCH_3$ to acetone. It is clear that uranium–carbon σ bonds constitute a new facet of organoactinide chemistry, and this area is under continuing exploration.[20]

Acknowledgments. We are grateful to the National Science Foundation (GP-30623X) and the Research Corporation for support of this work and to the staff of the Chemistry Department Analytical Services Laboratory for assistance with several measurements.

(9) Alternatively, but less plausibly, the spectral pattern is due to a highly distorted *trihapto*allyl structure.

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(16) (a) Decomposition of Ib in toluene-d_8 produces butane with only 5.0 ± 1.0% deuterium incorporation (determined mass spectrometrically[16c]). This suggests that hydrogen abstraction[18] occurs principally within a solvent cage from the cyclopentadienyl rings. Further studies are in progress. (b) Negligible reaction takes place between Ib and added 1-butene. (c) K. Biemann, "Mass Spectrometry-Organic Chemical Applications," McGraw-Hill, New York, N. Y., 1972, p 223.

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On Making Corrections for Donor Nonideality on Molecular Complex Equilibria

Sir:

The authors of a recent publication[1] questioned the validity of equilibrium constants evaluated by spectrophotometric means for the 1:1 association between a Lewis acid A and a Lewis base D to form the adduct AD (eq 1). They objected to expressing the equilib-

$$A + D \rightleftharpoons AD \qquad (1)$$

rium constant for eq 1 in terms of concentration, for the various species, instead of the activities. As evidence for their proposal that activities should be used instead of concentrations, they investigated the system caffeine–benzene in CCl_4 solvent using nmr for determining *K*, the equilibrium constant, and Δ_{AD}, defined as the difference in chemical shift between free and fully complexed caffeine.

Since Δ_{AD} is concentration independent, data evaluated to solve for Δ_{AD} and *K* should give the same value for Δ_{AD} regardless of the concentration scale used (*i.e.*, *molal* or *mole fraction*). Since they found that Δ_{AD} evaluated when the initial donor concentration, D^0, was expressed in molal units was significantly different than when D^0 was expressed in mole fraction units, they con-

(1) M. W. Hanna and D. G. Rose, *J. Amer. Chem. Soc.*, **94**, 2601 (1972).

cluded that activities must be used and not concentration units. They found Δ_{AD} to be the same when activities were used in molal or mole fraction units, but this is of no significance in itself because one is simply correcting back to the same reference (Henry's law) state.

Although we do not question the fact that Δ_{AD} must be concentration independent, we shall present arguments to show that the main difficulty with the reported system[1] involves the validity of using molal concentration for describing homogeneous chemical systems, such as in eq 1. Furthermore, we have previously reported[2] a system in which an activity correction of the base concentration destroyed a perfectly good fit of the data obtained using concentration units and concluded that the ratio of activity coefficients must be remaining constant for all of the species in the equilibrium. In an ion-pairing study which gave similar results, we concluded[3] contrary to Hanna and Rose[1] that one should not indiscriminately correct any one of the species in the equilibrium with activity coefficients, γ.

Since much work involving spectrophotometric studies of Lewis acid–base interaction involves the use of *molar* concentration for D^0 and A^0, we proceeded to solve the Hanna and Rose data for K and Δ_{AD} using molar units. Since our method[4,5] for the simultaneous determination of Δ_{AD} and K is much more rigorous than the one used by Hanna (Scatchard plots[6]), we also reevaluated his data for D^0 expressed in molal and mole fraction units. The results are shown in Table I.

Table I

Concn units	Equilibrium constant K[a,b]	Δ_{AD},[a,b] cps
Molal	0.311 ± 0.002 (0.0006)	53.3 ± 0.2 (0.05)
Mole fraction	1.04 ± 0.01 (0.001)	104.9 ± 0.9 (0.09)
Molar	1.083 ± 0.001 (0.00015)	98.2 ± 0.9 (0.1)

[a] Solved simultaneously by a computer-fit procedure (ref 4) similar to the one used for the evaluation of calorimetric data (ref 11) based on the Rose–Drago equation (ref 5). [b] Errors given are marginal standard deviations; those in parentheses are the conditional standard deviations (see ref 4).

As can be seen, satisfactory agreement exists between Δ_{AD} evaluated when D^0 is expressed in either molar or mole fraction units, but if molal units are used, a different value for Δ_{AD} is obtained. We feel molal units should not be used when describing chemical systems such as in eq 1 when spectrophotometric means are used for the investigation. In fact, a strong case has been presented concerning whether *molalities* have any physical basis at all to warrant their use. Molalities were primarily formulated to provide a concentration expression which is temperature independent and are usually used with regard to colligative properties.[7]

In spectrophotometric investigations on systems described by eq 1, one uses a microscopic model to interpret the experimental data and to obtain the association constant[1] and a structural parameter (Δ_{AD}). In order to have a consistent representation of the system, one should use "data" in consistent microscopic units. This idea has been developed by Scott in a recent publication.[8] Here, a simple quasi-chemical calculation for a lattice of donor, acceptor, and inert solvent (S) was used to obtain an expression relating the spectroscopic observable (Δ_{obsd}) to the mole fraction equilibrium constant and the corresponding *fractions* of donor (χ_D), solvent (χ_S), and acceptor (χ_A). Relative fractions are used as a direct consequence of the microscopic model which describes the system as a lattice of coordination number N, in which the sites are occupied by D, A, or S in their appropriate fractions. All pair interaction energies are the same except the one between D and A which gives rise to a perturbation in a spectroscopic observable (chemical shift for nmr). In order to use this microscopic model, one should use as concentration for D^0 and A^0 a unit appropriate to the lattice site fractions. Mole fraction is a macroscopic concentration scale which reflects the microscopic concentration. *Molarities* may be used to approximate the mole fraction if one is dealing with relatively dilute, ideal solutions of solutes and solvents. One is not justified in using molal concentration units for concentrated solutions when attempting to interpret spectrophotometric data because this unit does not accurately reflect the relative molecular site fractions of solute and solvent molecules.

This is clearly seen in the results for the caffeine–benzene system shown in Table I. The value for Δ_{AD} obtained when D^0 is expressed in mole fraction units is 104.9 cps. This is the correct value predicted by the model. Molar concentrations for D^0 result in a value of 98.2 cps which is close to but not exactly the true value. This discrepancy arises because of the deviations in these two units caused by the very high concentration of benzene in CCl_4. The donor concentration should not be pushed to the limits used in the benzene–caffeine system (~5.7 M). At lower concentrations, the molar concentration units may be used without much difficulty.

Strong support for our viewpoint can be obtained by comparing chemical shifts for chloroform–base systems evaluated by solving for K and Δ_{AD} by our procedure with Δ_{AD} values for 1:1 complexes reported in the literature.[9,10] In the latter studies, donors with large formation constants (hexamethylphosphoramide and ethyl acetate) are studied under conditions in which the chloroform is fully complexed. The Δ_{AD} values reported by these authors for these systems are in excellent agreement with those reported by us[4] from equilibrium constant studies using molar concentration units. For hexamethylphosphoramide values of 1.995 and 1.92 ± 0.04 ppm are found by Lichter and Roberts[9] and us, respectively. With ethyl acetate[10] values of 0.71 and 0.66 ± 0.06 ppm are found by these two procedures.

We want to emphasize that the equilibrium constants evaluated by these spectrophotometric experiments

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are "sociation" rather than "association" constants and the limitations brought forth by Scott[8] should be kept in mind when evaluating enthalpies of adduct formation by this procedure. Nevertheless, for systems having enthalpies of adduct formation greater than about 1.5 kcal/mol, these[8] difficulties should be minimal. The incorrect spectroscopic constants will be obtained only when $\gamma_{DA}/\gamma_D\gamma_A$ is varying or is a constant very much different than unity. To detect the former complication, we strongly advocate treating the data by a Rose–Drago procedure[4,5,11] and looking at the intersections as a function of concentration.

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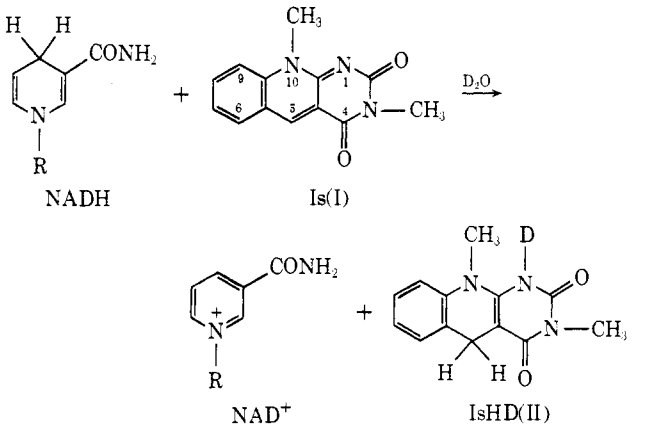
Demonstration of a Direct Hydrogen Transfer between NADH and a Deazaflavin

Sir:

Reductions by NADH have invariably been found to occur by direct transfer of a proton plus two electrons to the substrate molecule.[1] The mechanism of the biochemically important reduction of flavins (7,8-dimethyl-10-alkylisoalloxazines) by NADH is unknown though popular mechanisms reject direct hydride transfer and invoke covalent bond formation between the dihydronicotinamide and the flavin.[2] The difficulty in determining whether a direct hydride transfer occurs between NADH and flavin is undoubtedly due to the fact that the protons of the ultimate product (1,5-dihydroflavin) are bound to weakly basic nitrogens and are, therefore, exchangeable. Since the 5 nitrogen has been shown, *via* theoretical calculations,[3] to be the most electrophilic position of the flavin molecule and, therefore, the most likely recipient of a transferred hydride ion, we have investigated a compound where this nitrogen has been replaced by a carbon. The reaction of 3,10-dimethyl-5-deazaisoalloxazine (I)[4] with NADH has been examined in D_2O [80 mg of I suspended in 5 ml of D_2O containing 720 mg of the disodium salt of NADH was stirred for 3 days in the dark (argon atomsphere, 30°); the product (70 mg, 87%) was collected and washed with 2 ml of D_2O]. Except for the deuteron at position 1 the compound obtained was indistinguishable by nmr from that obtained on reduction of I with dithionite in H_2O, the nmr spectrum[5] establishing conclusively that the reduction product was IsHD(II) of the equation.



We have previously presented evidence that reduction of flavins by NADH takes place through a preequilibrium complex in which the dihdyronicotinamide does not occupy the area adjacent to the 1, 9, and 10 positions of the flavin.[6] If the present results are interpretable as evidence for transfer of two electrons from NADH to the flavin and a proton from the 4 position of the NADH to the 5 position of the flavin, then the geometry of the transition state becomes relatively defined. Evidence has been presented for direct hydrogen transfer from NADH to substrate *via* enzyme bound flavin.[7]

In passing it is of interest to note the general similarities of flavins and I. The second-order rate constants for the reactions of NADH and NPrNH with I [$k_{2,NADH}$ = 1.89 M^{-1} min^{-1} at pH 7.62, $k_{2,NPrNH}$ = 3.68 × 10^2 M^{-1} min^{-1} at pH 7.69 (30°, phosphate buffer containing 5 vol % DMF, μ = 0.19)] are not too dissimilar from the corresponding rate constants obtained with 3,10-dimethylisoalloxazine [$k_{2,NADH}$ = 53 M^{-1} min^{-1}, $k_{2,NPrNH}$ = 5.25 × 10^3 M^{-1} min^{-1}].[6] I reacts with the SO_3^{2-} component of sulfite buffer to provide the 5 adduct,[8] as previously shown for flavins.[9] Upon acidification of a sulfite–adduct solution in D_2O with DCl, pure I is generated as proven by the nmr spectrum. IsH_2 regenerates I on reaction with O_2,[10] as do 1,5-dihydroflavins,[11] and is oxidized by $(CH_3S)_2$ to yield I. The oxidation of mercaptans by flavins is well established[6,12,13] and the reduction of a disulfide by IsH_2 is the retrograde of this reaction. As in the case of flavins, I forms nonfluorescent complexes with tryptophan and β-resorcylic acid. The 1:1 complexing constants with tryptophan and β-resorcylic acid, determined by

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(4) *Anal.* Calcd for $C_{13}H_{11}N_3O_2$: C, 64.72; H, 4.60; N, 17.42. Found: C, 64.53; H, 4.55; N, 17.26.

(5) Nmr spectrum of I in $CDCl_3$: δ 8.95 [1, s, C(5)–H], 8.2–7.2 [4, m, Ar–H], 4.19 [3, s, N(10)–CH_3], 3.47 ppm [3, s, N(3)–CH_3]. Nmr spectrum of II in $CDCl_3$ (DMSO-d_6): δ 7.6–6.8 (7.6–6.8) [4, m, Ar–H], 3.83 (3.63) [2, s, C(5)–H_2], 3.51 (3.30) [3, s, N(10)–CH_3], 3.38 ppm (3.16) [3, s, N(3)–CH_3]. The proton at position 1 of IsH_2 shows a singlet at δ 4.49 ppm in $CDCl_3$.

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(8) 3,10-Dimethyl-5-sulfonate-5-deaza-1,5-dihydroisoalloxazine: λ_{max} 307 nm (ϵ 12,500 M^{-1} cm^{-1}) at pH 6.78 (10 vol % CH_3CN, μ = 0.9); K_{eq} = 3.88 × 10^2 M^{-1} at pH 6.93 (5 vol % DMF, μ = 0.19), 30°; k_f = 2.09 × 10^3 M^{-1} sec^{-1} at pH 6.89 (5 vol % DMF, μ = 0.1), 30°; nmr spectrum in D_2O δ 7.7–6.7 [4, m, Ar–H], 5.13 [1, s, C(5)–H], 3.33 ppm [6, s, N(3,10)–CH_3], the proton at position 1 shows a singlet at δ 5.4 ppm in H_2O.

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