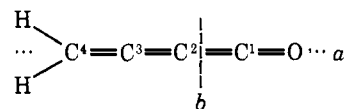**Table V.** Calculated Perpendicular Distance (d) of the Hydrogen Atoms from the a Axis

molecule	d from Kraitchman's equations/pm	d from I_a/pm
formaldehyde	96	94
ketene	94	94
butatrienone	100 ± 8	93 ± 6

We can also use Kraitchman's equations to compute the distance of the hydrogen atoms from the b axis, the value obtained being 314 ± 2 pm. Once again, we can derive from this the value of the HCH angle if we make the following assumptions, based on data[11] for ketene, allene, and butatriene, about bond lengths: C–O = 116 pm, C_1–C_2 = 130 pm, C_2–C_3

$$\begin{array}{c} H \\ \diagdown \\ \cdots \diagup C^4 \!\!=\!\! C^3 \!\!=\!\! C^2 \!\!=\!\! C^1 \!\!=\!\! O \cdots a \\ H \\ \big| \\ b \end{array}$$

= 129 pm, C_3–C_4 = 129 pm, C–H = 108 pm. The result is 121 $\pm 2.5°$. (The error limits were estimated by assuming that the adopted bond lengths may be uncertain by ± 1 pm.)

We conclude that, whereas the rotational spectrum of propadienone indicates nonlinearity of the heavy-atom chain, or

unusually large vibrational effects, the spectrum of butatrienone implies that the molecule is planar with C_{2v} symmetry, and, as far as we can tell, similar to formaldehyde and ketene in its general structural features.

References and Notes

(1) R. F. C. Brown, F. W. Eastwood, and G. L. McMullen, *J. Am. Chem. Soc.*, **98**, 7421 (1976).
(2) R. F. C. Brown, F. W. Eastwood, and G. L. McMullen, *Aust. J. Chem.*, **30**, 179 (1977).
(3) G. L. Blackman, R. D. Brown, R. F. C. Brown, F. W. Eastwood, and G. L. McMullen, *J. Mol. Spectrosc.*, **68**, 488 (1977).
(4) I. Heilbron, E. R. H. Jones, and F. Sondheimer, *J. Chem. Soc.*, 604 (1949).
(5) G. Eglinton, E. R. H. Jones, G. H. Mansfield, and M. C. Whiting, *J. Chem. Soc.*, 3197 (1954).
(6) W. D. Emmons, K. S. McCallum, and A. F. Ferris, *J. Am. Chem. Soc.*, **75**, 6047 (1953).
(7) J. S. Meunter, *J. Chem. Phys.*, **48**, 4544 (1968).
(8) G. W. King, R. M. Hainer, and P. C. Cross, *J. Chem. Phys.*, **11**, 27 (1943).
(9) W. Gordy and R. L. Cook, "Microwave Molecular Spectra", Interscience, New York, 1970: (a) pp 184–187; (b) pp 220–231.
(10) J. Kraitchman, *Am. J. Phys.*, **21**, 17 (1953).
(11) "Tables of Interatomic Distances and Configuration in Molecules and Ions. Supplement 1956–1959", *Chem. Soc., Spec. Publ.*, **No. 18** (1965).

Kinetic Studies on the Nucleophilic Addition to Double Bonds. 1. Addition of Amines to Electrophilic Carbon–Carbon Double Bonds[†]

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Abstract: Kinetic and thermodynamic data are reported for the primary elementary step of nucleophilic addition to polar carbon–carbon double bonds in aprotic solvents (acetonitrile and chloroform). As nucleophiles primary, secondary, and tertiary amines were applied. The rapid reaction techniques used are stopped-flow (SF) and temperature-jump (TJ) relaxation.

I. Introduction

The nucleophilic attack on carbon–carbon double bonds has been the subject of numerous investigations which were summarized in review articles.[1,2] It represents the primary step of nucleophilic addition and various other consecutive reactions. The former process, in general, follows a mechanism of type 1 which consists of two distinct elementary steps at least:

$$L + B \underset{k_{21}}{\overset{k_{12}}{\rightleftharpoons}} LB \tag{1a}$$

$$LB + E \underset{k_{32}}{\overset{k_{23}}{\rightleftharpoons}} P \tag{1b}$$

L stands here for the molecule containing the electrophilic carbon–carbon double bond. B is a base, e.g., an amine (B = RR'R″N). The primary addition step (1a) leads to an association complex (LB) which for electrically neutral B is a zwitterion. When B is a primary or secondary amine intramolecular

proton transfer may immediately follow reaction 1a or even occur simultaneously. The second reaction step (1b) involves an electrophile E and its nature will largely depend on the particular structure of L and B. There are various examples in which reaction 1a appears to be rate determining. In protic solvents usually the reaction with the proton (E = H^+) dominates. Commonly one works under conditions which make the step 1b irreversible for practical purposes ($k_{32} \sim 0$). The primary nucleophilic attack is often too fast for conventional kinetic studies and hence only equilibrium constants $K = k_{12}/k_{21}$ or overall rates $k' = k_{12}k_{23}/(k_{21} + k_{23})$ have been measured. Chemical relaxation and flow techniques offer a possibility to obtain direct information.

Stopped-flow (SF) measurements in most cases were found to be too slow to achieve the time resolution required. Temperature-jump (TJ) relaxation appears to be particularly suited as far as the time range is concerned but, on the other hand, has to face certain obstacles: in protic solvents the reaction mechanism is more involved than eq 1 indicates since both the solvent and its autodissociation product may act as nucleophiles as well and consequently several parallel reactions have to be included explicitly in a detailed kinetic analysis. Aprotic sol-

[†] Dedicated to Professor O. E. Polansky.

vents of low or medium polarity like chloroform, dioxane, or acetonitrile thus seem to be predestined for the desired investigations. The use of these solvents, however, creates another problem: in order to guarantee sufficiently fast Joule's heating rather large amounts of conductive salts have to be added. Various association equilibria may complicate the situation and careful choice of a suitable salt and systematic analysis of the equilibrium mixtures were found to be inevitable.

Direct relaxation studies on the primary addition step (1a) require k_{23} to be very small. This can be achieved by (1) absence of strong electrophiles, particularly protons, in the solution and/or (2) stabilization of the adduct LB by an appropriate choice of substituents. Powerful stabilization of the anion formed in the primary attack occurs in a class of compounds (L) which can be regarded as "cryptic" Lewis acids.[3-5] R_3 may

L
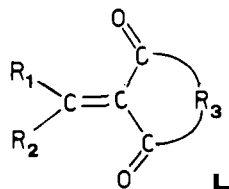
be any bidentate substituent which forces the two carbonyl groups into a coplanar or almost planar steric arrangement. The negative charge in the zwitterion LB then is efficiently delocalized within a planar extended π-electron system. In this

LB
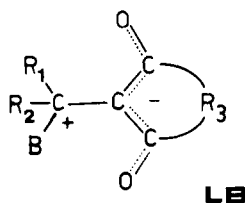
first paper of a series we present chemical relaxation and stopped-flow studies on the addition of some amines, namely, *n*-butylamine, benzylamine, piperidine, morpholine, diethylamine, *N*-methylpiperidine, *N*-methylmorpholine, to compounds of class L, in particular to substituted benzylidene-Meldrum's acids (BMS)[6] and –*N,N'*-dimethylbarbituric acids (BBS). Acetonitrile and chloroform were chosen as aprotic

BMS   **BBS**
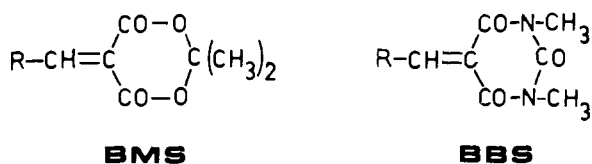
solvents. In the first case the dielectric constant is sufficiently high in order to allow for temperature-jump relaxation studies by Joule heating.

II. Materials and Methods

Compounds and Purification. Benzylidene–Meldrum's acid (BMS), benzylidene-*N,N'*-dimethylbarbituric acid (BBS), and their para-substituted derivatives were prepared and purified according to procedures described in the literature.[7,8] The nitrogen bases were freshly distilled from CaH_2 under nitrogen atmosphere. Tetraalkylammonium salts were recrystallized from ethanol. Acetonitrile was a high-purity commercial material (Merck, Uvasol) used without further purification. Chloroform of the same quality (Merck, Uvasol stabilized with traces of ethanol) was applied. Removal of the stabilizer did not change the measurements within the limits of error.

Equilibrium constants were determined spectroscopically. The extinctions of solutions of the benzylidene compounds with different base concentrations were measured on a Zeiss PMQ II spectrophotometer. The temperature (25 °C) was controlled directly in the sample cell. The data obtained thereby were evaluated according to Benesi and Hildebrand.[9] The good straight lines found in these plots (Figure 1) confirm 1:1 stoichiometry; only one reaction product is
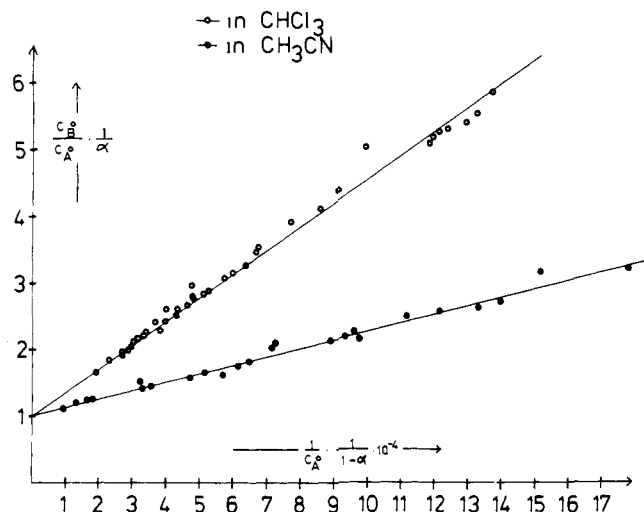
Figure 1. Determination of the equilibrium constant K of the system BMS/piperidine in chloroform and acetonitrile according to Hildebrand and Benesi[9] (α, the degree of dissociation, is obtained from the UV spectroscopic measurement of extinctions: $\alpha = (E - E_0)/(E_\infty - E_0)$ wherein E_0 refers to the solution with initial base concentration $c_B{}^0 = 0$ and E_∞ to complete reaction (vanishing free concentration of A: $c_A \sim 0$ or lim $c_B{}^0 \rightarrow \infty$)).

therefore possible, spectroscopically identified as the addition complex LB.[3,4] At high base concentrations[10] ($c_B{}^0 > 0.1$ M) the extinctions of the solutions become time dependent and consecutive reactions leading to irreversible decomposition take place.

Temperature-Jump (TJ) Relaxation Studies. TJ experiments were performed using a commercial spectrophotometer with Joule heating (Messanlagen Studiengesellschaft, Göttingen). Typical conditions applied throughout this paper follow: initial temperature $T_i = 23$ °C, final temperature $T_f = 25$ °C, conductive salt $(CH_3)_4N^+BF_4^-$, $c_S{}^0 = 0.05$ M, benzylidene derivative (BMS, BBS, etc.) 6×10^{-5} M $\leqslant c_A{}^0 \leqslant 1 \times 10^{-4}$ M, and base 5×10^{-5} M $\leqslant c_B{}^0 \leqslant 5 \times 10^{-2}$ M. Under these conditions the time for heating the solutions amounts to $\tau_H \sim 10$ μs, thus being by far shorter than the shortest chemical relaxation time measured here.

In the case of the system BMS/piperidine additional measurements were performed at final temperatures $T_f = 17, 20, 30$, and 35 °C in order to determine reaction and activation enthalpies and entropies, respectively.

Stopped-flow (SF) experiments were performed on a commercial stopped-flow spectrometer (Durrum, Model D 110), $T = 25$ °C.

Recording and Processing of Data. Two procedures for evaluation of relaxation times were applied.

(1) The output data of the spectrometer were sampled and stored in digital form on a transient recorder (Biomation, Type 805), then transferred to a tape and via a data terminal (Texas Instruments, Silent 700 ASR) to a CDC Cyber 73 computer. Relaxation times were obtained by a numerical fit of an exponential function to the samples recorded.

(2) The relaxation curve stored was compared directly with an exponential function of known and variable time constant by a simple analogue simulator which consists of a resistance–capacity integrating circuit.[11]

Procedure 1 was generally applied for the evaluation of whole series of experiments. The relaxation times obtained in this way were fitted to eq 2, which represents the analytical expression for a single step association reaction:

$$\tau^{-1} = k_{12}(\bar{c}_A + \bar{c}_B + K^{-1}) \tag{2}$$

$\bar{c}_A$ and $\bar{c}_B$ are equilibrium concentrations of A and B, respectively. The constants k_{12} and K thus were determined simultaneously from kinetic data (Table I). The final plots—see, e.g., Figure 2—are fully consistent with eq 2, and show relatively little scatter and no systematic deviation. In the system *p*-nitro-BBS/piperidine the validity of the single-step approach has been checked independently by fitting the measured relaxation amplitudes to the function Γ_c given by the
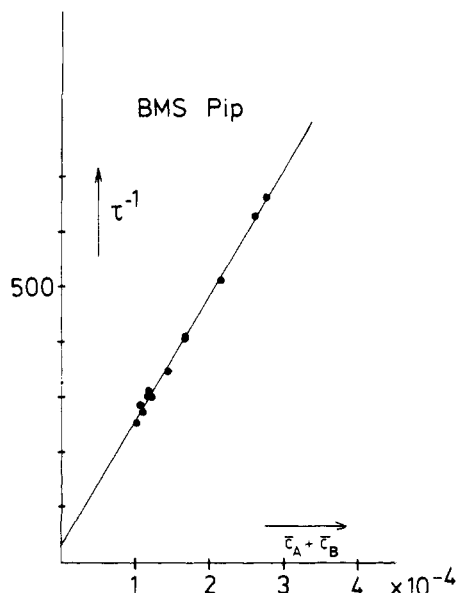
Figure 2. Typical plot of the reciprocal relaxation time against the sum of the equilibrium concentrations (BMS/piperidine in acetonitrile).

Figure 3. Relaxation amplitudes of the reaction of p-NO_2-BBS with piperidine in acetonitrile. A function Γ $(c_A{}^0, c_B{}^0, K)$[14,15] as explicitly given in eq 3 was fitted to the experimental points $(c_A{}^0 = 8 \times 10^{-5}, T_f = 25\,°C, \Delta T = 2\,°C, K = 1.15 \times 10^6\,M^{-1})$.

Table I. Rate and Equilibrium Constants for the Primary Addition of Piperidine to BMS Obtained by Different Techniques (Solvent Acetonitrile, $T = 25\,°C$)

method	k_{12}, $M^{-1}\,s^{-1}$	k_{21}, s^{-1}	K, M^{-1}
UV spectroscopy			7.40×10^4
stopped flow	2.16×10^6	32	6.77×10^4
temperature jump[a]	2.37×10^6	33	7.25×10^4
	2.16×10^6	29	7.45×10^4

[a] Conductive salt: 0.05 M $(CH_3)_4N^+BF_4^-$ corresponding to $I \sim 0.02$ M (the ion pair association constant was assumed to be similar to that of $(CH_3)_4N^+I^-$ [12,13]). The results of two independent series of experiments are given.

equation[14,15]

$$\Gamma_c = \left(\frac{\partial c_A}{\partial \ln K}\right)_{P,T}$$

$$= \frac{1}{2K} \frac{\bar{c}_A + \bar{c}_B + K^{-1}}{(\bar{c}_A - \bar{c}_B)^2 + 2K^{-1}(\bar{c}_A + \bar{c}_B) + K^{-2}} - 1 \quad (3)$$

The value of the equilibrium constant obtained thereby agrees well with that determined by conventional spectrophotometric investigation (Figure 3).

Stopped-flow experiments were evaluated with respect to both initial slopes (lim (dc_A/dt), $t \rightarrow 0$) and relaxation conditions (evaluation of the linear range around equilibrium: $\Delta c_A(t) \leq 0.05\Delta c_A(0)$, $\Delta c_A = |c_A - \bar{c}_A|$).

A comparison of rate and equilibrium constants obtained from different experiments is given in Table I. As we can see, the scatter of data from different techniques, although it is less than 10%, is larger than or of the same order of magnitude as the influence of ionic strength in the temperature-jump (TJ) experiments ($I \sim 0.02$ M).

Conductive Salts. The use of tetraethyl- and tetrabutylammonium salts provided a certain, unexpected difficulty. These salts contained a hard to remove small contamination ($\leq 0.5\%$) which occurred independently of the anion—ClO_4^-, BF_4^-, Cl^-, Br^-, and I^- have been used. The impurity has been identified[16] as the salt of the corresponding tertiary amines which presumably is formed by a Hoffmann-type elimination. The salt of a tertiary amine exchanges a proton with the base B applied in reaction 1 and thereby disturbs the thermodynamic and kinetic studies. The effect is particularly strong when the two amines differ largely with respect to their proton affinities (e.g., B = piperidine).
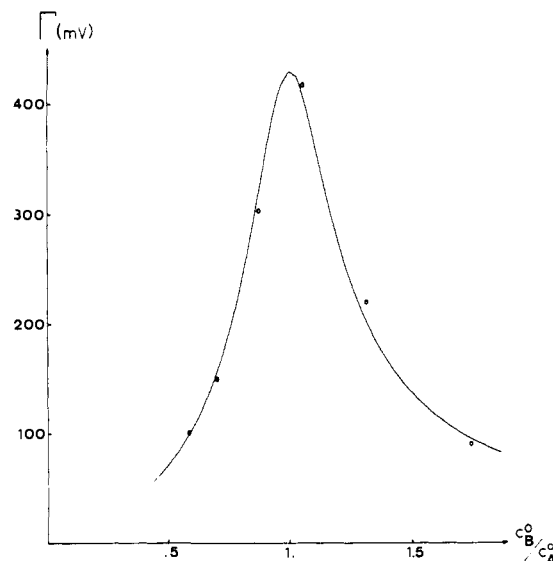
Table II. Comparison of Rate and Equilibrium Constants for the Primary Addition of Amines to the Electrophilic Double Bond of BMS in Chloroform and Acetonitrile ($T = 25\,°C$)

solvent and method[b]	amine[a]	k_{12}, $M^{-1}\,s^{-1}$	k_{21}, s^{-1}	K, M^{-1}
CH_3CN, TJ	PIP	2.3×10^6	32	7.4×10^4
	MOR	4.0×10^5	310	1.3×10^3
$CHCl_3$, SF	PIP	1.2×10^6	57	2.2×10^4
	MOR	1×10^5	700	1.4×10^2

[a] PIP = piperidine; MOR = morpholine. [b] TJ = temperature jump relaxation, $I = 0.02$ M, conductive salt $(CH_3)_4N^+BF_4^-$. SF = stopped flow.

III. Results and Discussion

Rate and equilibrium constants for the primary addition step (1a) were determined according to section II and are summarized in Tables II, III, and V. It is appropriate to split the discussion into five parts.

Effects of Solvent and Conductive Salts. The two solvents applied in this study differ mainly with respect to their dielectric constants ($CHCl_3$, $\epsilon = 4.7$; CH_3CN, $\epsilon = 36$[13]). As we can see from Table II the association constants K are reduced on changing the solvent from acetonitrile to chloroform. Thus the association complex (LB) is less stable in chloroform, the solvent with the lower ϵ value. This result agrees with a polar zwitterionic structure of LB. In the BMS/piperidine system the bimolecular rate constant k_{12} is reduced by a factor of 2 whereas the first-order dissociation rate constant k_{21} increases by about the same factor on going from acetonitrile to chloroform. Thus the two rate constants are changed in opposite direction and accordingly there is a marked solvent effect on the equilibrium constant. With morpholine as base we observe essentially the same changes, although the differences in the constants are somewhat larger and the influence on k_{12} is more pronounced.

No drastic influence of the ionic strength on reaction 1a was observed in agreement with other studies on related reactions.[17] This is not unexpected since both reactants (L and B) and the association complex (LB) are electrically neutral species and hence there will be no primary salt effect.

Influence of Amine Structure and Basicity. Rate and equilibrium constants for the primary addition step depend strongly

Table III. Comparison of Rate and Equilibrium Constants for the Primary Addition of Amines to Benzylidene-N,N'-Dimethylbarbituric Acid (BBS) in Acetonitrile (T = 25 °C)

	amine	method[a]	k_{12}, M^{-1} s^{-1}	k_{21}, s^{-1}	K, M^{-1}
primary	*n*-butylamine	TJ, SF	3.5×10^4	0.12	2.9×10^5
	benzylamine	SF	1.3×10^4	0.7	1.7×10^4
	piperidine	TJ	6.8×10^5	2.7	2.6×10^5
secondary	morpholine	TJ	1.5×10^5	20	7.5×10^3
	diethylamine	TJ	7.9×10^4	20	4.0×10^3
	quinuclidine				35
tertiary	*N*-methylpiperidine	TJ	180	18	10
	N-methylmorpholine	TJ	130	42	3.1

[a] TJ = temperature jump relaxation; SF = stopped flow.

on the nature of the amine (Table III). Both the basicity of the amine and stereochemical factors are of importance. The influence of base strength is seen best from a comparison of the data obtained for piperidine and morpholine, which have almost identical stereochemistry. The stronger base piperidine associates faster by a factor of 4.5. The great difference in equilibrium constants, nevertheless, is brought about by an even larger difference in complex dissociation constants: k_{21} for piperidine is smaller by a factor of 6.5 in comparison to the morpholine adduct. Tertiary amines, as we see already at a superficial glance at Table III, associate with BBS much weaker than primary or secondary amines do—the association constants are smaller by more than two orders of magnitude. The difference in complex stability may be explained by hydrogen-bond formation (L·NRR'H) which cannot occur with tertiary amines.

L·NRR'H
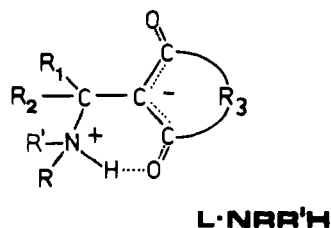
A similar type of intramolecular hydrogen bond has been postulated in order to explain kinetic and thermodynamic data for amine addition to ω-nitrostyrene,[18] and even earlier to interpret the stereochemistry of morpholine addition to a properly activated carbon-carbon double bond[1] and in a discussion of solvent effects on the reactions of secondary amines with diethyl 2-chloro-2-*p*-nitrophenylethylene-1,1-dicarboxylate.[19] This interpretation will be supported somewhat by reaction and activation entropies reported later. It seems important to note that both rate constants change when going from secondary to tertiary amines, but it is the association rate k_{12} which accounts for the major part of the difference in complex stability. Experiments with deuterated piperidine -ND, at present in progress in our laboratory, presumably will provide more information. An alternative explanation of the enormous gap in association rate constants k_{12} between secondary and tertiary amines is provided by steric hindrance.

Stereochemical factors, indeed, play an extremely important role as a comparison of the results for piperidine and diethylamine shows. "Bending back" the arms through ring closure leads to an increase in k_{12} and a simultaneous decrease in k_{21} by almost one order of magnitude. Thus the equilibrium constant K is larger by a factor of about 60 in the system BBS/ piperidine. There is no pronounced difference in the rate and equilibrium constants between primary and secondary amines.

There is a great variety of literature on various kinetic studies of nucleophilic addition to polar double bonds and aromatic systems. Only a few investigations, however, provide

Table IV. Comparison of Association Rate Constants (k_{12}) for Nucleophilic Addition of Amines to Carbon–Carbon Double Bonds and Aromatic Systems

	association rate constant k_{12}, M^{-1} s^{-1} for electrophile	
amine	BBS	2,4-dinitrophenyl phenyl ether
piperidine	6.8×10^5	3.1×10^6
morpholine	1.5×10^5	3.3×10^5
n-butylamine	3.5×10^4	1.5×10^5
solvent	acetonitrile	dioxane/water (10/90)
T, °C	25	29.4
ref	this work	20

reliable information on the rate constants of the primary nucleophilic attack. For the purpose of comparison we choose as an example a recent study in which conventional kinetic methods were used to calculate association rate constants (Table IV). The sequence in reactivity of the amines investigated appears to be the same in both reactions. A comparison of the rate constants with respect to their absolute values does not seem to be meaningful since they refer to different media.

Comparison of BMS and BBS and the Influence of Para Substitution on the Benzene Ring. Benzylidene-N,N'-dimethylbarbituric acid (BBS) and its derivatives have a higher affinity toward piperidine than the corresponding derivatives of benzylidene-Meldrum's acid (BMS), although the latter associate faster by a factor of about 3 (Table V). Thus the difference in equilibrium constants which amounts to a factor of approximately 5 is related to the much slower dissociation rate of the adduct of BBS and its derivatives. Evidently, the predominant difference in behavior has to do with the relative stability of the two complexes BMS/piperidine and BBS/ piperidine. Greater stability of the second complex may have its origin in a stronger hydrogen bond of the type discussed above (L·NRR'H) or in better stabilization of the anionic part of the complex by the solvent. It is interesting to note in this context that both compounds, BMS and BBS, have roughly the same "Lewis acidity" in methanol ($pK_{BMS} \sim pK_{BBS} = 9.2$, T = 25 °C, solvent CH$_3$OH).[8,21] On the basis of the presently available information it is hard to distinguish between the two explanations. Experiments with deuterated amine presumably will clarify this question.

Rate and equilibrium constants of both series, BMS and BBS derivatives, fulfill linear free energy (Hammett) relations provided that σ^+ values[22] are applied (Table V). The ρ values derived fit well into the long list reported by Rappoport and Ladkani.[2] Lough and Currie[18] reported conventional kinetic studies on the addition of morpholine to ω-nitrostyrene and derivatives in dioxane, a reaction which comes closest to that reported here. The ρ value for their calculated rate constants ($\rho = 0.69^2$) agrees well with our results (BMS, $\rho = 0.79$; BBS,

Table V. Comparison of Rate and Equilibrium Constants for the Primary Addition of Piperidine to Substituted Benzylidene-Meldrum's Acids (BMS) and Benzylidene-N,N'-Dimethylbarbituric Acids (BBS) in Acetonitrile (T = 25 °C)[a]

substituent (R)	σ^+ Hammett value[17]	BMS k_{12}, M⁻¹ s⁻¹	BMS k_{21}, s⁻¹	BMS K, M⁻¹	BBS k_{12}, M⁻¹ s⁻¹	BBS k_{21}, s⁻¹	BBS K, M⁻¹
p-(CH₃)₂N-	-1.71	1.1×10^5	580	1.9×10^2	2.9×10^4	29	1.0×10^3
p-CH₃O-	-0.77	9.5×10^5	110	8.4×10^3	3.2×10^5	7.6	4.2×10^4
H-	0.	2.3×10^6	32	7.4×10^4	6.8×10^5	2.7	2.6×10^5
p-Br-	+0.15	3.6×10^6	30	1.2×10^5	1.1×10^6	2.3	4.5×10^5
p-NO₂-	+0.79				1.6×10^6	1.2	1.3×10^6
linear regression[b] ρ		0.79	-0.70	1.49	0.70	-0.59	1.26
Δ		0.07	0.03	0.04	-0.01	0.04	-0.07
r^2		0.980	0.995	0.992	0.938	0.995	0.972

[a] Rate constants were obtained from temperature-jump (TJ) relaxation studies (I = 0.02 M, conductive salt $(CH_3)_4N^+BF_4^-$). [b] Linear regression analysis of the Hammett equation: $\log k_R - \log k_H = \rho\sigma^+(+\Delta)$, r^2 represents the regression coefficient. Thus r^2 = 1 and Δ = 0 mean perfect correlation.

Table VI. Rate and Reaction Quantities Obtained from Kinetic and Equilibrium Studies of the Addition of Piperidine to BMS in Acetonitrile at Different Temperatures (T = 17, 20, 25, 30, and 35 °C)[a]

system	quantity measured	reaction or activation enthalpy, kJ/mol	reaction or activation entropy, J/mol·K	regression[b] coefficient r^2	activation energy, kJ/mol	regression[b] coefficient r^2
BMS +	k_{12}	16.7	-66.9	0.806	19.2	0.848
piperidine	k_{21}	88.7	80.8	0.990	91.2	0.990
	K	-71.1	-146.4	0.998		
2,4-dinitro-	k_{12}	-6.7	-101.7		-4.2	
phenol +	k_{21}	56.5	40.6		59.0	
tri-n-octylamine	K	-57.3	-121.3			

[a] For comparison the analogous values of the reaction of tri-n-octylamine with 2,4-dinitrophenol are given as well.[24,25] [b] Linear regression $y = a + bx$ to n points; $r^2 = [\Sigma x_i y_i - (\Sigma x_i \Sigma y_i/n)]^2/[\Sigma x_i^2 - ((\Sigma x_i)^2/n)(\Sigma y_i^2 - ((\Sigma y_i)^2/n)]$.
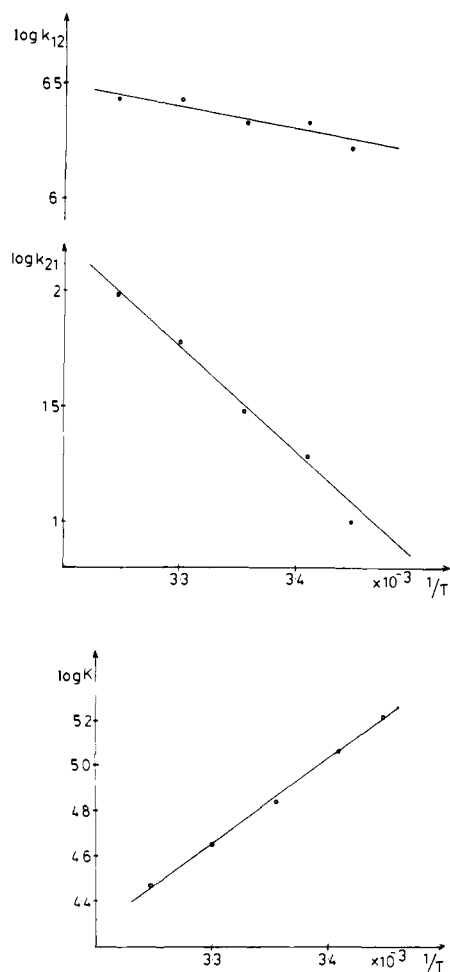

Figure 4. Temperature dependence of equilibrium and rate constants in the system BMS/piperidine.

ρ = 0.70). As far as the equilibrium constants are concerned our ρ values lie very close to that reported by Rappoport and Gertler[23] for the reaction of tributylphosphine to benzylidenemalodinitrile and derivatives (T = 34 °C, ρ = 1.27). The dissociation rate of the complex is characterized by a negative ρ value. Thus, electron-withdrawing substituents stabilize the association complex by two ways: they make the benzylidene derivative react faster and slow down the complexes dissociation rates. The substituent and solvents effects (Tables II and V) determined in the system BMS/piperidine suggest the assumption that the transition state of the reaction is roughly in the middle between reactants and products.

Reaction and Activation Enthalpies and Entropies. In order to learn more about the nature of reaction 1a we determined rate and reaction quantities from a study of the temperature dependence in the system BMS/piperidine. The rate constant for the dissociation process as well as the equilibrium constant shows the expected linear correlations in $\ln k/T^{-1}$ or $\ln K/T^{-1}$ plots (Figure 4). The temperature dependence of the association rate constant k_{12} is much weaker and hence the scatter of data is larger. Nevertheless, a small positive activation enthalpy or energy with a significance $r^2 > 0.8$ was obtained also in the latter case (Table VI).

The addition of piperidine to the double bond in BMS is characterized by negative activation and reaction entropies. This is not unexpected since the formation of an encounter complex and of a defined aggregate usually is accompanied by a loss in entropy. The large difference between $\Delta S^{\ddagger}_{12}$ and ΔS shows that the reaction complex loses a substantial amount of entropy after it passes the transition state. The first explanation at hand again invokes hydrogen bond formation subsequent to the addition of a secondary or primary amine which makes the whole complex a rigid entity due to ring closure (L·NRR'H). Alternatively, the formation of a highly polar structure like this zwitterion itself leads to structural reordering in the surrounding solvent and thus may be the cause of highly negative reaction entropies as well.

Comparison of the Reactions of Lewis and Brønsted Acids with Amines. It seems interesting to compare the addition of an amine to a kind of Lewis acid studied here with the analogous Brønsted acid–base reaction. Superficially proton transfer between phenols and amines seems to have little in common with reaction 1a. When we consider, however, that the ion pairs formed do hardly dissociate in media of low dielectric constant such as in chlorobenzene, which was used in this study, we observe a very similar sequence of reaction steps:[24,25]

$$AH + B \rightleftharpoons AH,B \rightleftharpoons AH \cdots B \rightleftharpoons A^- \cdots HB^+ \qquad (4)$$
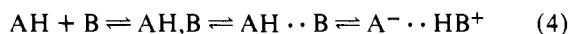
Again a highly polar complex, here an ion pair, is formed by a primary addition step and reordering of the encounter complex. For comparison we give the reaction and activation quantities of an example of this type of reaction, the complex formation between 2,4-dinitrophenol and tri-*n*-octylamine, in Table VI.[25] Similar to our example large negative values for reaction and activation entropies are reported.

So far we have treated reaction 1a as a single elementary step, not controlled by diffusion, which is consistent with the kinetic data (see section II) and the positive value of $\Delta H^{\ddagger}{}_{12}$ or $E_a{}^{\ddagger}$, respectively. Complex formation between phenols and amines studied in chlorobenzene behaves somewhat differently[24,25] (Table VI): in this case formally negative values of $\Delta H^{\ddagger}{}_{12}$ and $E_a{}^{\ddagger}$ were obtained which demand a mechanism of at least two steps in order to explain the negative sign by a preequilibrium.

References and Notes

(1) S. Patai and Z. Rappoport in "The Chemistry of Alkenes", S. Patai, Ed., Interscience, New York, 1964, pp 469–583.
(2) Z. Rappoport and D. Ladkani, *Chem. Scr.*, **5**, 124–133 (1974).
(3) P. Schuster, O. E. Polansky, and F. Wessely, *Tetrahedron*, **8**, 463–483 (1966).
(4) F. Kunz, P. Margaretha, and O. E. Polansky, *Chimia*, **24**, 165–181 (1970).
(5) H. McNab, *Chem. Soc. Rev.*, **7**, 345 (1978).
(6) Meldrum's acid is the convenient shorthand name for 2,2-dimethyl-1,3-dioxane-4,6-dione.[5]
(7) P. Schuster, O. E. Polansky, and F. Wessely, *Monatsh. Chem.*, **95**, 53 (1964).
(8) R. Bednar, E. Haslinger, U. Herzig, O. E. Polansky, and P. Wolschann, *Monatsh. Chem.*, **107**. 1115 (1976).
(9) H. A. Benesi and J. H. Hildebrand, *J. Am. Chem. Soc.*, **71**, 2703 (1949).
(10) Throughout this paper we use the symbol $c_X{}^0$ for initial and c_X for free concentrations. A, B, and S refer to compound L, the base B, and the conductive salt: L = c_A, B = c_B, etc.
(11) J. E. Crooks, M. S. Zetter, and P. A. Tregloan, *J. Phys. E.*, **3**, 73 (1970).
(12) U. Mayer, *Coord. Chem. Rev.*, **21**, 159 (1976).
(13) G. J. Janz and R. P. Tomkins, "Non-Aqueous Electrolytes Handbook", Vol. 1, Academic Press, New York, 1972, p 562.
(14) R. Winkler-Oswatitsch, Doctoral Thesis, Technical University of Vienna and Max-Planck-Institute for Physical Chemistry, Göttingen, 1969.
(15) M. Eigen and R. Winkler-Oswatitsch, *Angew. Chem.*, **91**, 20 (1979).
(16) B. Schreiber, M. Hoffmann-Ostenhof, P. Wolschann, and P. Schuster, submitted for publication.
(17) R. Ta-Shma and Z. Rappoport, *J. Am. Chem. Soc.*, **98**, 8460 (1976).
(18) C. E. Lough and D. J. Currie, *Can. J. Chem.*, **44**, 1563 (1966).
(19) Z. Rappoport and A. Topol, *J. Chem. Soc., Perkin Trans. 2*, 863 (1975).
(20) C. F. Bernasconi, R. H. de Rossi, and P. Schmid, *J. Am. Chem. Soc.*, **99**, 4090 (1977).
(21) P. Schuster, A. Stephen, O. E. Polansky, and F. Wessely, *Monatsh. Chem.*, **99**, 1246 (1968).
(22) H. C. Brown and Y. Okamoto, *J. Am. Chem. Soc.*, **80**. 4979 (1958).
(23) Z. Rappoport and S. Gertler, *J. Chem. Soc.*, 1360 (1964).
(24) E. F. Caldin, J. E. Crooks, and D. O'Donnell, *J. Chem. Soc. Faraday Trans. 1*, **69**, 993 (1973).
(25) E. F. Caldin and K. Tortschanoff, *J. Chem. Soc. Faraday Trans. 1*, **74**, 1804 (1978).

Comparative Chemistry of the Bay- and Non-Bay-Region Tetrahydro Epoxides of Phenanthrene

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Abstract: It has been previously proposed that the great mutagenic and carcinogenic activity of the tetrahydro epoxides of a number of polycyclic aromatic hydrocarbons, in which the epoxide moiety resides at a "bay region", is due to the greater stability of the bay region benzylic carbonium ion formed upon oxirane ring opening. The rate constants for specific-acid-, general-acid-, and water-catalyzed oxirane ring opening of the bay (I) and nonbay (II) epoxides of 1,2,3,4-tetrahydrophenanthrene are compared. Also compared are the second-order rate constants for nucleophilic attack upon I and II. With few notable exceptions the various second-order rate constants for the various reactions of I and II were found to be quite similar with the rate constants for I exceeding those of II by an approximate average of threefold. Similar rate ratios are seen in the reactions of the bay and nonbay phenanthrene arene oxides wherein the bay benzylic carbonium ion stability cannot be a factor. A notable exception is the general-acid-catalyzed oxirane ring opening by H_3PO_4, where the second-order rate constant for I exceeds that for II by 40-fold. Brønsted α (-0.7) and β_{nuc} ($+0.25$) values obtained with I and II are indistinguishable. Also similar are the values of $\Delta S^{\ddagger}$ and kinetic solvent isotope effects (k^{H_2O}/k^{D_2O}) for H_2O- and H_3O^+-catalyzed oxirane ring opening. The values of $\Delta S^{\ddagger}$ and k^{H_2O}/k^{D_2O} for H_3O^+-mediated solvolysis are as expected for a specific acid mechanism. Nevertheless, the values of the log $k_{H_3O^+}$ fit accurately to the Brønsted plots for general acid catalysis. Product studies show a mixture of cis and trans diols to be formed at low and intermediate pHs, while only trans diols are formed at high pH. These results are discussed.

Introduction

There is presently strong evidence that the ultimate carcinogenic metabolites of benzo[a]pyrene and benz[a]anthracene are dihydrodiol epoxides in which the epoxide function is in the bay region[1] of the saturated angular benzo ring.[2a-g] On the basis of perturbational molecular orbital calculations it has been suggested that bay-region epoxides should be more re-

active than corresponding non-bay-region epoxides owing to the greater calculated stability of the benzylic bay region carbonium ion.[3] Studies have confirmed that the bay-region dihydrodiol epoxides of benzo[a]pyrene,[2g] benz[a]anthracene,[2a,4] and chrysene[5] are more highly mutagenic than are the corresponding non-bay-region dihydrodiol epoxides. However, it has not been shown that bay and nonbay epoxides