Thermodynamic Analysis of Proton Dissociation Equilibria of Some 4-Diethylaminoazobenzenes in 20 wt% Aqueous Methanol Solution

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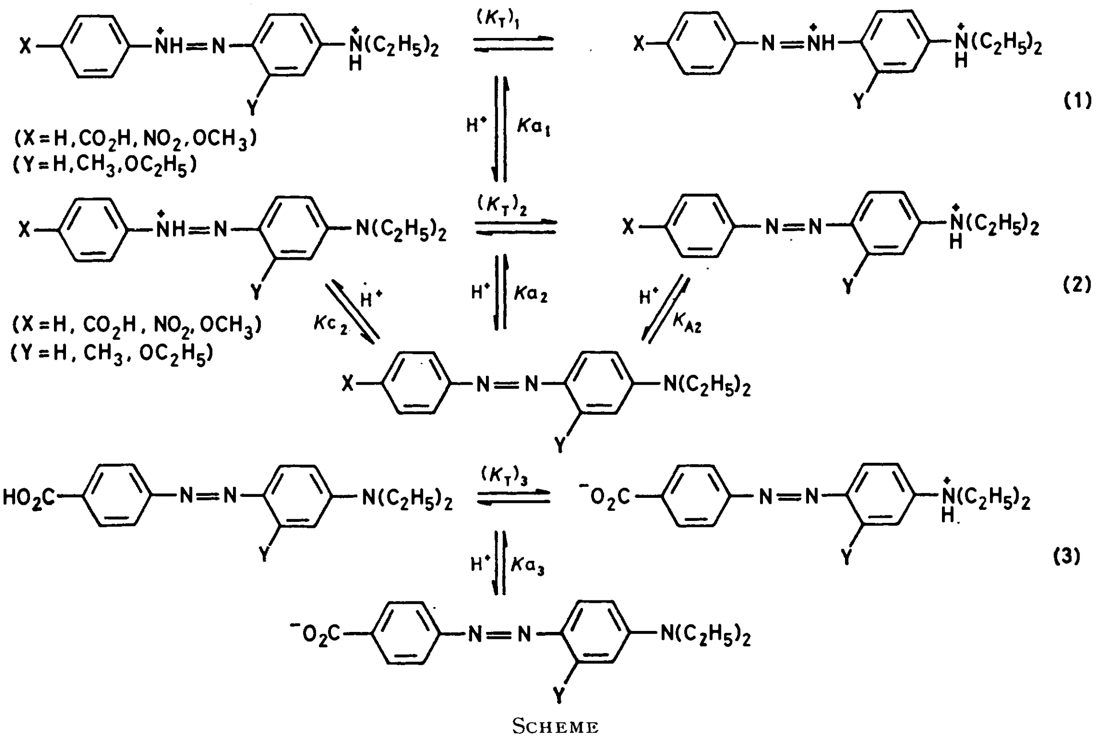
Proton dissociation of $XC_6H_4\overset{+}{N}C{=}NH_6H_3Y\overset{+}{N}HEt_2$ (X = H, CO_2H, NO_2, OMe; Y = H, Me, OEt) has been examined at 293.15 K in aqueous sulphuric acid solutions. Equilibrium constants for the proton dissociation of $XC_6H_4\overset{+}{N}H{=}NC_6H_3YNEt_2$ (X = H, CO_2H, NO_2, OMe; Y = H, Me, OEt) have been measured as a function of temperature, in the range 293—323 K, in 20 wt% aqueous methanol solution. Thermodynamic functions ($\Delta G°$, $\Delta H°$, $\Delta S°$) have been calculated. Experimental data have been discussed and interpreted in terms of linear free energy relationships. Tautomeric equilibrium constants have been evaluated and apparent equilibrium constants separated into the separate contributions of ammonium and azonium forms.

THE acid–base equilibria of substituted azobenzenes have been the subject of much discussion.[1-3] The problem of tautomerism between the ammonium and azonium forms of 4-dialkylaminoazobenzenes has been particularly examined.[4-8]

Being interested in the photochemical behaviour of 4-dialkylaminoazobenzenes in different solvents, in aqueous alcoholic solvents specifically, and as the photochemistry of these substrates is deeply influenced by solvent effects and acid–base properties, we deemed it useful to make a thorough thermodynamic analysis relative to some molecules which were considered to be suitable models for this study. To tackle the problem, we have chosen, in this paper, the unsubstituted compound and its 4′-methoxy, 4′-carboxy, 4′-nitro, 4′-carboxy-2-ethoxy, and 4′-carboxy-2-methyl derivatives.

We have investigated the acid–base equilibria of reactions (1)—(3) (see Scheme, where K_T represents tautomeric equilibrium constants, and K_a thermodynamic constants of proton dissociation equilibria), with the purpose of rationalizing K_T in terms of acidity and substituent effects, as well as of dividing apparent equilibrium constants into the separate contributions of the ammonium and azonium forms.



SCHEME

EXPERIMENTAL

Materials.—Spectrograde solvents were used throughout the experiments. All the products were kindly supplied by ACNA S.p.A. and purified by repeated crystallizations from Bu^nOH.

pK_a *Measurements.*—pK_a Measurements were carried out spectrophotometrically according to a standard technique.[9] For the study of equilibria (2) and (3) the solvent was 20 wt% aqueous methanol. In this solvent [10] the difference between the measured pH and pa_{H}^{*} was considered to be −0.06 pH units at 298.15 K. Buffers at different pa_{H}^{*} and at constant ionic strength (I 0.045M) were prepared by mixing in appropriate ratios aqueous alcoholic solutions (i) and (ii)

TABLE 1

pK_a (K_a/mol l^{-1}) values for proton dissociation reactions (1)—(3) as a function of temperature

	X = H, Y = H		X = CO_2H, Y = H			X = NO_2, Y = H	
T/K	$-pK_{a_1}$ [a,b]	pK_{a_2} [c]	$-pK_{a_1}$ [a,b]	pK_{a_2} [d]	pK_{a_3} [e]	$-pK_{a_1}$ [a,b]	pK_{a_2} [f]
293.15	3.42 ± 0.05	2.75 ± 0.02	3.30 ± 0.05	2.66 ± 0.01	5.74 ± 0.01	5.74 ± 0.05	1.84 ± 0.01
298.15		2.70 ± 0.02		2.61 ± 0.02	5.71 ± 0.01		1.82 ± 0.02
303.15		2.64 ± 0 02		2.57 ± 0.01	5.70 ± 0.01		1.80 ± 0.01
308.15		2.58 ± 0.01		2.53 ± 0.02	5.70 ± 0.02		1.78 ± 0.01
313.15		2.53 ± 0.04		2.49 ± 0.01	5.73 ± 0.02		1.76 ± 0.01
318.15		2.48 ± 0.02		2.45 ± 0.02	5.71 ± 0.01		1.74 ± 0.03
323.15		2.46 ± 0.01		2.41 ± 0.03	5.72 ± 0.02		1.72 ± 0.05

	X = OCH_3, Y = H		X = CO_2H, Y = CH_3			X = CO_2H, Y = OC_2H_5		
T/K	$-pK_{a_1}$ [g,b]	pK_{a2} [h]	$-pK_{a_1}$ [a,b]	pK_{a_2} [b]	pK_{a_3} [i]	$-pK_{a_1}$ [a,e]	pK_{a_2} [l]	pK_{a_3} [l]
293.15	1.86 ± 0.05	2.99 ± 0.01	4.9 ± 0.1	3.13 ± 0.01	6.06 ± 0.06	5.3 ± 0.1	4.41 ± 0.02	5.64 ± 0.02
298.15		2.94 ± 0.02		3.07 ± 0.01	6.08 ± 0.01		4.42 ± 0.09	5.63 ± 0.05
303.15		2.90 ± 0.04		3.02 ± 0.01	6.11 ± 0.05		4.38 ± 0.08	5.60 ± 0.04
308.15		2.86 ± 0.02		2.97 ± 0.01	6.06 ± 0.06		4.35 ± 0.09	5.67 ± 0.02
313.15		2.81 ± 0.01		2.91 ± 0.03	6.13 ± 0.05		4.34 ± 0.09	5.66 ± 0.03
318.15		2.77 ± 0.01		2.9 ± 0.1	6.21 ± 0.03		4.32 ± 0.09	5.64 ± 0.02
323.15		2.73 ± 0.03		2.83 ± 0.05	6.39 ± 0.03		4.30 ± 0.09	5.62 ± 0.02

[a] In aqueous sulphuric acid. [b] Wavelength range for spectroscopic determination, 530—510 nm. [c] Wavelength range for spectroscopic determination, 540—520 nm. [d] Wavelength range for spectroscopic determination, 550—530 nm [e] Wavelength range for spectroscopic determination, 500—480 nm. [f] Wavelength range for spectroscopic determination, 470—450 nm. [g] Wavelength range for spectroscopic determination, 490—470 nm. [h] Wavelength range for spectroscopic determination, 460—440 and 550—540 nm. [i] Wavelength range for spectroscopic determination, 430—410 nm. [l] Wavelength range for spectroscopic determination, 520—500 nm.

[(*i*) 0.025M-NaOAc–0.025M-KCl; (*ii*) 0.05M-KCl–0.1M-HCl] with an aqueous methanol solution of the substrate. Mean ionic activity coefficients $\gamma\pm$ were estimated from the Debye–Hückel approximation (4) where D is dielectric constant of the solvent, T is absolute temperature, and a_i is the distance of closest approach of ions (taken as 4.5 Å). Ionic activity coefficients of uncharged species were assumed to be unitary. For pK_{a_1} determination [equilibrium (1)], aqueous sulphuric acid solutions (up to 96 wt% H_2SO_4) were used. H_0 Values were taken from the literature.[11,12] The concentration ratio was estimated spectrophotometrically by absorbance measurements in solutions in which the equilibrium under study could be considered displaced to the right or to the left ($pa_H^* \gtrless pK_a \pm 2$), as well as in 6—8 buffer solutions of known pa_H^*, all absorbances being measured under strictly thermostatted conditions. The linearity of the absorbance–concentration plot for each species was verified in a wide concentration range (1×10^{-5}—5×10^{-3}M). Spectroscopic measurements were made with an Optica model 10 spectrophotometer equipped with a thermostatted cell compartment. The experimental ranges are given in Table 1. pK_a Values of proton dissociation equilibria and their standard deviations were obtained by a least-squares linear regression analysis.

$$-\log\gamma\pm = \frac{1.812\,3 \times 10^6\, I^{\frac{1}{2}}}{(D\,T)^{\frac{3}{2}}[1 + 50.288 \times 10^8\, (D\,T)^{-\frac{1}{2}} a_i\, I^{\frac{1}{2}}]} \quad (4)$$

RESULTS AND DISCUSSION

Equilibrium (1) was measured at 293.15 K in aqueous sulphuric acid solutions. Equilibria (2) and (3) were investigated in 20 wt% aqueous methanol solutions as a function of temperature in the range 293—323 K. Determination of acid–base concentration ratio was carried out spectrophotometrically for equilibria (2) and (3) in buffered solutions at constant ionic strength (0.045M). Spectral data of all possible species are reported in Table 2.

The results in the form of pK_a values and their standard deviations are reported in Table 1. Plots of pK_a against $1/T$ gave a reasonable straight line for all the compounds examined. This behaviour showed that the assumption of temperature-independent enthalpy ($\Delta H°$) and entropy ($\Delta S°$) for equilibria (2) and (3) was generally valid in the temperature range investigated. Thus, the best that could be done was to obtain mean $\Delta H°$ and $\Delta S°$ values in that range. For the calculation of thermodynamic functions from equilibrium data, $\Delta H°$ was

TABLE 2

Molar absorptivities (ε l mol^{-1} cm^{-1}) at maximum absorption wavelengths ($\lambda_{max.}$/nm) of 4-diethylaminoazobenzenes and their conjugate acids in 20 wt% aqueous methanol solution

	X = H, Y = H		X = CO_2H Y = H		X = NO_2, Y = H		X = OCH_3, Y = H		X = CO_2H, Y = CH_3		X = CO_2H, Y = OC_2H_5	
	$\lambda_{max.}$	log ε	$\lambda_{max.}$	log ε	$\lambda_{max.}$	log ε	$\lambda_{max.}$	log ε	$\lambda_{max.}$	logε	$\lambda_{max.}$	log ε
$XC_6H_4\overset{+}{N}H{=}NC_6H_3Y\overset{+}{N}HEt_2$	416 [a]	4.70 [a]	413 [a]	4.65 [a]	411 [a]	4.54 [a]	470 [a]	4.84 [a]	431 [a]	4.68 [a]	465 [a] 384 [a]	4.37 [a] 4.41 [a]
$XC_6H_4N{=}NC_6H_3Y\overset{+}{N}HEt_2$	317	4.40 [b]	319	4.51 [b]	322	4.61 [b]	354	4.45 [b]	289	4.13 [b]	274	4.97 [b]
$XC_6H_4\overset{+}{N}H{=}NC_6H_3YNEt_2$	516	4.62 [b]	513	4.85 [b]	510	4.82 [b]	550	4.34 [b]	513	4.82 [b]	485	4.86 [b]
$XC_6H_4N{=}NC_6H_3YNEt_2$	423	4.49	463	4.56	495	4.56	443	4.49	474	4.57	480	4.57
$^-O_2CC_6H_4N{=}NC_6H_3YNEt_2$			447	4.59					444	4.54	463	4.58

[a] In concentrated sulphuric acid solution. [b] Values extrapolated according to the method of ref. 17.

TABLE 3

Thermodynamic functions for proton dissociation reactions (2) and (3)

	X = H, Y = H	X = CO_2H, Y = H		X = NO_2, Y = H	X = OCH_3, Y = H	X = CO_2H, Y = CH_3		X = CO_2H, Y = OC_2H_5	
	Reaction (2)	Reaction (2)	Reaction (3)	Reaction (2)	Reaction (2)	Reaction (2)	Reaction (3)	Reaction (2)	Reaction (3)
$\Delta G°_{293.15}$/ kJ mol^{-1}	15.35 $\pm$0.13	14.93 $\pm$0.13	32.60 $\pm$0.13	10.38 $\pm$0.13	16.82 $\pm$0.13	17.53 $\pm$0.13	35.23 $\pm$0.13	25.10 $\pm$0.13	32.17 $\pm$0.13
$\Delta H°_{298.15}$/ kJ mol^{-1}	18.3 $\pm$1.1	14.6 $\pm$1.2	0.4 $\pm$1.2	7.8 $\pm$1.1	15.5 $\pm$1.2	18.0 $\pm$1.1	0.4 $\pm1.3^{a}$	7.0 $\pm$1.2	-0.4 $\pm$1.3
$\Delta S°_{298.15}$/ J mol^{-1} K^{-1}	10 $\pm$4	0 $\pm$ 4	-109 ± 4	-8 ± 4	-4 ± 4	0 $\pm$ 4	-117 ± 4^{a}	-59 ± 4	-109 ± 4

[a] In the temperature range 308—323 K a marked deviation from linearity of log K_{a_2} *versus* $1/T$ plot for this compound was observed. In this range mean $\Delta H°$ and $\Delta S°$ are, respectively, -38.5 ± 1.5 and -242 ± 5.

obtained by a linear regression analysis of a log K_a *versus* $1/T$ plot and $\Delta S°$ by the linear function of the free energy changes of reactions (2) and (3) against the temperature. The thermodynamic data are collected in Table 3. The entropy values of reaction (3) parallel the behaviour of substituted benzoic acids [13] and point out generally that the solvated conjugate acids, independently of the position of their zwitterionic equilibrium (see Scheme), are less ordered than the solvated bases. For reaction (2), in contrast, there is no marked entropy effect, other than for 4'-carboxy-2-ethoxy-4-diethylaminoazobenzene. The moderately low entropy value for reaction (2) relative to this compound may indicate partially at least, its greater zwitterionic contribution.

Experimental pK_{a_2} values, relative to the acid ionization equilibrium (2), refer to the apparent equilibrium constant, because, as is evident from the Scheme, in these conditions we always deal with a mixture of tautomers. The problem of determining $(K_T)_2$ is complicated by the fact that it is impossible to separate the two tautomeric species. To overcome this difficulty, several authors [14–16] characterize the position of the equilibrium by the intensity ratio of the molar absorptivity maximum (ε_A) of band A at *ca.* 300 nm (attributed to the ammonium form) and that (ε_C) of band C at *ca.* 510 nm (ascribed to the azonium form). Gerson and Heilbronner [6] assumed, on the basis of the data by Isaks and Jaffé,[5] $K_T \simeq 3\ \varepsilon_A/\varepsilon_C$ for some substituted dimethylaminoazobenzenes. The approximations involved in these assumptions are rather crude and have been the subject of much criticism by various authors.[8] Even if a rough idea of the tautomeric equilibrium can be deduced in this way, absolute conclusions on K_T values cannot be drawn unambiguously by this procedure. The fact that tautomeric equilibrium shifts with change in the acid concentration enabled Bershtein and Ginzburg [17] to develop a method for determining absolute K_T values. The application of this latter method to our compounds was successful, while a constant ratio $(K_T)_2\varepsilon_C/\varepsilon_A$, as postulated by Isaks and Jaffé,[5] is not generally observed (see Table 4). Determination of the tautomeric equilibrium constants $(K_T)_2$ from the absorption spectra of our substrates at different acid concentrations has been carried out based upon the fact that the apparent molar extinction coefficient of band A decreases while simultaneously that of band C increases with increase in acidity. Values of log $(K_T)_2$ varied linearly with the Hammet acidity function H_o of the solutions, following equation (5). Parameters for equation (5) for the compounds investigated are collected in Table 4, together with $(K_T)_2$ values at H_o -1.0. The variation of K_{T_2} with acidity arises from the difference in the protonation behaviour between amino and azo sites: their values (see Table 2), in fact, parallel the difference between pK_{a_2} and pK_{a_1}. The increase in acidity invariably corresponds to a shift of tautomeric equilibrium towards the azonium form, as a consequence of a greater stability of this form in the more acid media, because of solvation effects.[17,18] The marked influence of substituents both on $(K_T)_2$ values, which are reflected by parameter b of equation (5) and on their variation with acidity [parameter a of equation (5)] represent quantitatively this behaviour.

$$\log K_T = aH_o + b \qquad (5)$$

TABLE 4

Values of $(K_T)_2$ (see Scheme) (at H_o -1.0 and T 293.15 K), parameters of equation (5), and $(K_T)_2\ \varepsilon_C/\varepsilon_A$ ratios for $XC_6H_4N{=}NC_6H_3YNEt_2$

Substituents	a	b	$(K_T)_2$	$(K_T)_2\ \varepsilon_C/\varepsilon_A$
X = H, Y = H	0.213	0.554	2.19	1.65
X = CO_2H, Y = H	0.348	0.273	0.84	2.25
X = NO_2, Y = H	0.144	-0.275	0.38	1.59
X = OCH_3, Y = H	0.946	1.357	2.58	0.77
X = CO_2H, Y = CH_3	0.234	-0.814	0.09	4.9
X = CO_2H, Y = OC_2H_5	0.051	-0.817	0.42	0.76

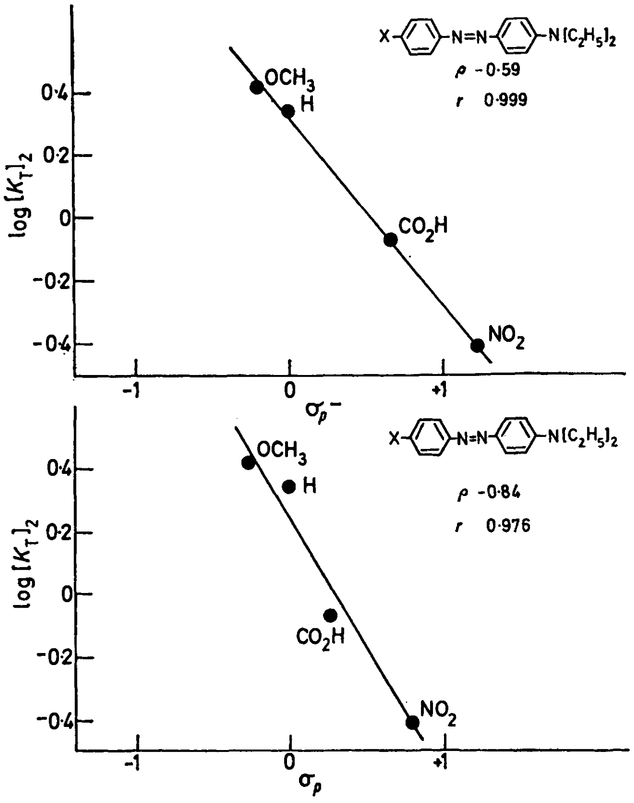
The correlation of log$(K_T)_2$ values by the σ–ρ plot with σ_p [19] for 4'substituents works fairly well (ρ values and correlation coefficients r are given in the Figure), but is much improved by the use of σ_p^-,[20] which stresses the dominant role of conjugation in the effect of 4'-substituents on the position of tautomeric equilibrium constants. Furthermore, a knowledge of the $(K_T)_2$ values allows a calculation of the microconstants K_{C_2} and K_{A_2} be-

TABLE 5

pK_{A_2} and pK_{C_2} values for reaction (2) at 298.15 K

	pK_{A_2}	pK_{C_2}
X = CO_2H, Y = H	2.47	2.20
X = CO_2H, Y = CH_3	2.25	3.06
X = CO_2H, Y = OC_2H_5	3.53	4.35

tween the azonium or ammonium forms, respectively, and the deprotonated form (see Scheme). The logarithms of K_{G_2} and K_{A_2}, the values of which are reported in Table 5, for the 2-substituted derivatives when plotted against σ_o [21] and σ_m [19] yield values of 5.53 (r 0.999) and −6.96 (r 0.978) respectively. The negative ρ value for the K_{A_2} plot indicates a prevailing role for the zwitterionic form in this equilibrium, a conclusion which seems perfectly in line with the entropy variation of reaction (2). The K_{A_2} microconstant has a slight effect on the composition of the macroconstant K_{a_2}, as $(K_T)_2$ values for 2-substituted 4′-carboxy-derivatives are shifted towards the azonium form. This fact is also shown by the satisfactory linearity of the corresponding log K_{a_2} *versus* σ_o plot (ρ 4.55, r 0.982), the azo group of the azonium structure being the reaction centre. In any case, the separate treatment of the contributions of the ammonium and azonium forms to the acid–base properties of 4-diethylaminoazobenzenes rationalizes the substituent and medium effects satisfactorily.



log $(K_T)_2$ *versus* substituent constants (σ_p or σ_p^-) for tautomeric equilibrium (2) (see Scheme). ρ Values and correlation coefficients r are indicated. Temperature 293.15 K; H_0 −1.0

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REFERENCES

[1] A. Hantzsch and F. Hilscher, *Ber.*, 1917, **50**, 856.

[2] A. Hantzsch and A. Buroway, *Ber.*, 1930, **63**, 1760.

[3] G. E. Lewis, *Tetrahedron*, 1960, **10**, 129.

[4] I. Ya. Bershtein and V. I. Minkin, *Zhur. org. Khim.*, 1969, **5**, 2207.

[5] Si. Jung Yeh and H. H. Jaffé, *J. Amer. Chem. Soc.*, 1959, **81**, 3283; M. Isaks and H. H. Jaffé, *ibid.*, 1964, **86**, 2209.

[6] F. Gerson and E. Heilbronner, *Helv. Chim. Acta*, 1962, **45**, 42.

[7] A. Zenhaüsern and H. Zollinger, *Helv. Chim. Acta*, 1962, **45**, 1882.

[8] I. Ya. Bershtein and O. F. Ginzburg, *Russ. Chem. Rev.*, 1972, **41**, 97.

[9] I. R. Bellobono and P. Beltrame, *J. Chem. Soc.* (*B*), 1969, 620; I. R. Bellobono and G. Favini, *ibid.*, 1971, 2034.

[10] C. L. Deligny, P. F. M. Luykx, M. Renbach, and A. A. Wienecke, *Rec. Trav. chim.*, 1960, **79**, 699, 727.

[11] M. J. Jorgensen and D. R. Hartter, *J. Amer. Chem. Soc.*, 1963, **85**, 878.

[12] J. C. Brand, *J. Chem. Soc.*, 1950, 997.

[13] J. W. Larson and L. G. Hepler, in 'Solvent–Solute Interactions,' eds. J. F. Coetzee and C. D. Ritchie, Dekker, New York, 1969, p. 11.

[14] W. C. J. Ross and G. P. Warwick, *J. Chem. Soc.*, 1956, 1719.

[15] I. N. Zhmurova, *Zhur. obshchei Khim.*, 1957, **27**, 2704.

[16] E. Sawicki, *J. Org. Chem.*, 1956, **21**, 605.

[17] I. Ya. Bershtein and O. F. Ginzburg, *Doklady Akad. Nauk. S.S.S.R.*, 1968, **98**, 180.

[18] S. Yamamoto, N. Nishimura, and G. Hasegawa, *Bull. Chem. Soc. Japan*, 1973, **46**, 194; S. Yamamoto, Y. Tenno, and N. Nishimura, *Austral. J. Chem.*, 1979, **32**, 41.

[19] H. H. Jaffé, *Chem. Rev.*, 1953, **53**, 992.

[20] C. D. Ritchie and W. F. Sager, *Progr. Phys. Org. Chem.*, 1964, **2**, 323.

[21] R. W. Taft, *J. Amer. Chem. Soc.*, 1952, **74**, 3120.