Reaction of *NN'*-Dimethylimidazolidine-2-thione with Methyl Iodide. Kinetic and Thermodynamic Aspects

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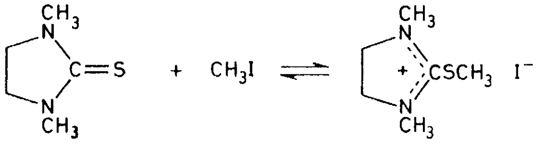
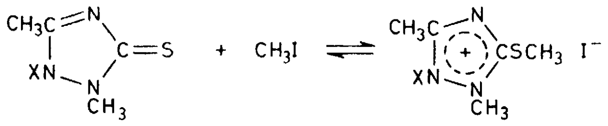
The reaction of *NN'*-dimethylimidazolidine-2-thione with methyl iodide followed simple second-order kinetics in polar solvents, whereas in less polar solvents the reverse process made a significant contribution to the overall rate. The solvent effects on the forward rate constant were linearly correlated with those on the rate of the reaction of tetramethylthiourea with methyl iodide. However, the reactions of two 1,2,4-triazole-3-thione derivatives with methyl iodide did not show any such correlations. From the volume changes and enthalpy changes measured, three transition state indices, $(\Delta_2V^{\ddagger}/\Delta_2\dot{V}^{\circ})$, n_T, and α were determined in acetonitrile, *ca.* 0.28, 0.35, and 0.43, respectively.

PROGRESS has recently been made in the rationalisation of solvent effects on kinetic and thermodynamic phenomena both in terms of multiple-parameter correlations [1-5] and in terms of theoretical calculations.[6] As a result of this work it seems necessary to study any given process by as many different methods as possible. We have therefore used free energy, enthalpy, and volume relationships to study a number of reactions between electrically neutral species.

In this report we analyse the reaction of *NN'*-dimethylimidazolidine-2-thione with methyl iodide which has a convenient rate[7] from both kinetic and thermodynamic aspects, and then for comparison we discuss the rates of the reaction of 2,3-dihydro-1,2,5-trimethyl- and 2,3-dihydro-2,5-dimethyl-1-phenyl-1,2,4-triazole-3-thione with methyl iodide.


(DMIT) (MDMIT$^+$ I$^-$)


X = CH_3 (TTT)
X = C_6H_5 (DPTT)

EXPERIMENTAL

Materials.—*NN'*-Dimethylimidazolidine-2-thione, prepared as described,[8] was recrystallized four times from water and was then dried *in vacuo* over phosphorus pentaoxide at 65 °C. 2,3-Dihydro-2,5-dimethyl-1-phenyl-1,2,4-triazole-3-thione, prepared as described,[9] was recrystallized three times from ethanol and dried *in vacuo* over phosphorus pentaoxide at 65 °C. 2,3-Dihydro-1,2,5-trimethyl-1,2,4-triazole-3-thione was prepared from *NN'*-dimethylhydrazine monohydrochloride by essentially the same procedure as described, except that instead of pouring the mixture of thiosemicarbazide, acetic anhydride, and acetic acid into water, it was poured into ether (Found: C, 41.7; H, 6.4; N, 29.15; S, 22.1. Calc. for $C_5H_9N_3S$: C, 41.95; H, 6.35; N, 29.35; S, 22.4%). 1,3-Dimethyl-2-methylthioimidazolidinium iodide, prepared by refluxing the corresponding thione and methyl iodide in acetonitrile, was recrystallized three times from acetonitrile–ether and was dried *in vacuo* over phosphorus pentaoxide at 65 °C (Found: C, 26.8; H, 4.75; N, 10.65; S, 24.75. Calc. for $C_6H_{13}IN_2S$: C, 26.5; H, 4.8; N, 10.3; S, 24.6%).

Solvents were purified as described elsewhere.[10 11] Kinetic measurements and density measurements were performed as described elsewhere.[10, 11]

The apparent molal volumes were calculated by equation (1) where the quantities have their usual meaning.

$$\phi_V = 1\,000\,(d_0 - d)/cd_0 + M_2/d_0 \qquad (1)$$

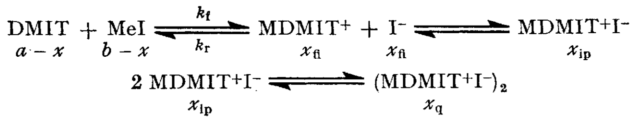
Heat of Reaction Measurements.—The heat of reaction was measured at 25 °C with a Tokyo Riko MPC-11 calorimeter used as a conduction type vessel. A methyl iodide solution (0.200M) in acetonitrile or in propylene carbonate (25 ml) and dimethylimidazolidine-2-thione solution (0.1204 g dissolved in 2 ml of the respective solvent) sealed in an ampoule were placed in the calorimeter cell. After thermal equilibration, the ampoule was broken with stirring and the temperature change was recorded on a chart over 22 h. Calibration was achieved by introducing a measured quantity of electric energy. Two to three determinations yielded a heat of reaction ΔH in acetonitrile of -11.9 ± 0.2 and in propylene carbonate of -9.4 ± 0.2 kcal mol^{-1}.

Heat of Solution Measurements.—Heats of solution were measured at 25.0 °C with a twin isothermal calorimeter (Tokyo Riko TIC-2D) at 1/100M for salt, 3/100M for *NN'*-dimethylimidazolidine-2-thione, and 1/10M for methyl iodide. The experimental error was estimated to be *ca.* 1% from duplicate or triplicate runs.

RESULTS AND DISCUSSION

Rate Laws.—In polar solvents the reaction of *NN'*-dimethylimidazolidine-2-thione with methyl iodide followed second-order kinetics as expected,[7,12] *i.e.*, first order in both *NN'*-dimethylimidazolidine-2-thione and methyl iodide. In cyclohexanone, bromobenzene, and benzene, an equilibrium was eventually reached. For a comprehensive treatment of the kinetic data the Scheme was postulated, leading to equations (2)—(7). The subscript e denotes the concentration at equilibrium.

In cyclohexanone the stoicheiometric equilibrium constant defined by equation (8) is not a true constant



$$\underset{a-x}{\mathrm{DMIT}} + \underset{b-x}{\mathrm{MeI}} \underset{k_r}{\overset{k_f}{\rightleftharpoons}} \underset{x_{fi}}{\mathrm{MDMIT^+}} + \underset{x_{fi}}{\mathrm{I^-}} \rightleftharpoons \underset{x_{ip}}{\mathrm{MDMIT^+I^-}}$$

$$2\,\underset{x_{ip}}{\mathrm{MDMIT^+I^-}} \rightleftharpoons \underset{x_q}{\mathrm{(MDMIT^+I^-)_2}}$$

SCHEME

but varies according to the values of x_e. Combination of equations (3) and (6)—(8) gives (9).

From the set of x_e and K_{sto} values the equations were

$$\mathrm{d}x/\mathrm{d}t = k_f(a - x)(b - x) - k_r x_{fi}^2 \qquad (2)$$

$$K = k_f/k_r \qquad (3)$$

$$K_{ip} = x_{ip}/x_{fi}^2 \qquad (4)$$

$$K_q = x_q/x_{ip}^2 \qquad (5)$$

$$\begin{aligned} x &= x_{fi} + x_{ip} + 2\,x_q \\ &= x_{fi} + K_{ip}x_{fi}^2 + 2K_qK_{ip}^2x_{fi}^4 \end{aligned} \qquad (6)$$

$$k_f(a - x_e)(b - x_e) = k_r x_{fi,e}^2 \qquad (7)$$

$$K_{sto} = x_e^2/[(a - x_e)(b - x_e)] \qquad (8)$$

$$(K_{sto}/K)^{\frac{1}{2}} = 1 + K_{ip}(Kx_e^2/K_{sto})^{\frac{1}{2}} + 2K_qK_{ip}^2(K\,x_e^2/K_{sto})^{\frac{3}{2}} \qquad (9)$$

solved simultaneously to give K 1.35 × 10^{-3}, K_{ip} 2.08 × 10^3 l $\mathrm{mol^{-1}}$, and K_q 11.2 l $\mathrm{mol^{-1}}$. All these equilibrium constants for the present reaction seem to be over a reasonable range compared with previous results.[12] The forward rate constant k_f was calculated from integration of equation (10) by the procedure described.[12]

$$\int_0^x \frac{\mathrm{d}x}{(a - x)(b - x) - x_{fi}^2/K} = \int_0^t k_f \mathrm{d}t \qquad (10)$$

In bromobenzene and benzene, the ions are expected to exist as higher aggregates [13,14] and the assumption $x_q \gg x_{fi}$ and x_{ip} seems plausible. Under these conditions, the observed equilibrium constant as defined by the left-hand side of equation (11) leads to the right-hand side of equation (11). A forward rate constant k_f

$$K_{obs} \equiv x_e^{\frac{1}{2}}/[(a - x_e)(b - x_e)] = K(2K_qK_{ip}^2)^{\frac{1}{2}} \qquad (11)$$

was obtained by performing the integration as described [12] and comparison of the experimental sets of

$$\int_0^x \frac{\mathrm{d}x}{(a - x)(b - x) - \frac{1}{2}K_{obs}^{-1}} = \int_0^t k_f\,\mathrm{d}t \qquad (12)$$

x *versus* t data with the theoretical values calculated on the basis of the Scheme indicates fair agreement (Figure 2).

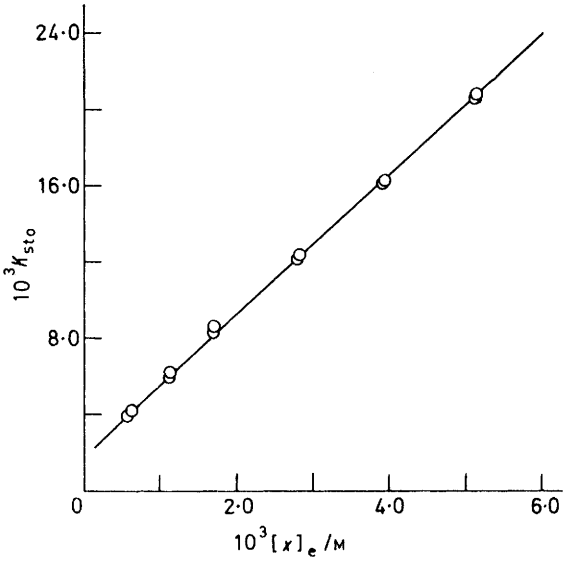


FIGURE 1 Plots of K_{sto} *versus* x_e at 30 °C, ○, experimental; ——, calculated with K 1.35 × 10^{-3}, K_{ip} 2.08 × 10^3 l $\mathrm{mol^{-1}}$, and K_q 11.2 l $\mathrm{mol^{-1}}$

According to the Scheme, the observed equilibrium constant should be a true constant. In fact there is internal agreement on a fixed value for a given initial concentration. The analytically determined equilibrium constants K_{obs} (16.6 in bromobenzene for $a = b =$ 4.00 × 10^{-2}M and 3.98 $\mathrm{l^{3/2}\,mol^{-3/2}}$ in benzene for $a = b =$ 8.00 × 10^{-2}M) agree well with the kinetically determined K_{obs} values (18.0 and 3.95 $\mathrm{l^{3/2}\,mol^{-3/2}}$, respectively) and also the calculated values for x *vs.* t data well reproduced the experimental values. In spite of this, there still exists an inconsistency in the Scheme, since the equilibrium constant K_{obs} showed a sudden increase at $a = b$ *ca.* 2.0 × 10^{-2}M with decreasing initial concentration in bromobenzene and, in contrast, at $a = b$ *ca.* 5.0 × 10^{-2}M with increasing initial concentration in benzene solvent.

In low dielectric solvents such as bromobenzene and benzene, highly polar substrates would solvate the un-

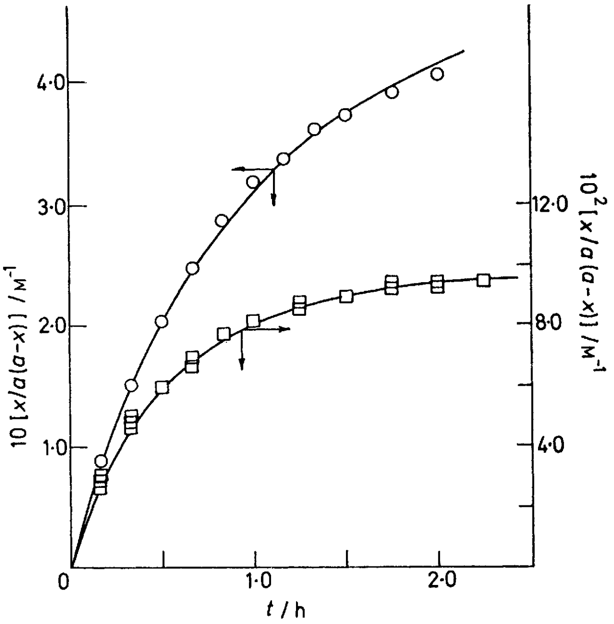


FIGURE 2 Plots of $[x/a\ (a - x)]$ *versus* t at 30 °C. In bromobenzene: ○, experimental; ——, calculated with K_{obs} 18.0 $\mathrm{l^{3/2}\,mol^{-3/2}}$, k_f 2.09 × 10^{-4} l $\mathrm{mol^{-1}\,s^{-1}}$, $a = b =$ 4.00 × 10^{-2}M. In benzene: □, experimental; ——, calculated with K_{obs} 3.95 $\mathrm{l^{3/2}\,mol^{-3/2}}$, k_f 7.80 × 10^{-5} l $\mathrm{mol^{-1}\,s^{-1}}$, $a = b =$ 8.00 × 10^{-2}M

stable ion aggregates more specifically than the bulk solvent itself does, thus making the kinetic behaviour more complex.

In contrast to the above example, the reactions of two triazole-3-thione derivatives TTT and DPTT with methyl iodide followed essentially second-order kinetics even in cyclohexanone solvent. This is the first instance of this phenomenon which we have met for thione derivatives. The sole equilibrium constant obtained (*K ca.* 170 for 2,3-dihydro-1,2,5-trimethyl-1,2,4-triazole-3-thione with methyl iodide in cyclohexanone) indicates that the equilibrium lies much further to the right compared with the reaction of *NN'*-dimethylimidazolidine-2-thione with methyl iodide (K 1.35 × 10^{-3}).

TABLE 1

Comparison of the rate constants at 30 °C

	DMIT + MeI		TTT + MeI		DPTT + MeI	
Solvents	$10^4\ k_f$/l mol^{-1} s^{-1}	$10^3\ k_x$/s^{-1}	$10^2\ k_f$/l mol^{-1} s^{-1}	$10\ k_x$/s^{-1}	$10^2\ k_f$/l mol^{-1} s^{-1}	$10\ k_x$/s^{-1}
Propylene carbonate	21.2	24.9	25.3	29.5	13.8	16.2
Acetonitrile	7.75	14.6	11.3	21.3	6.13	11.5
Nitrobenzene	9.74	9.44	26.4	25.5	13.6	13.2
Cyclohexanone	7.21	6.89	44.3	42.3	18.0	17.2
Bromobenzene	2.09	1.97				
Benzene	0.780	0.867				
Methanol	1.70	4.16	0.576	1.41	0.451	1.10
Propan-2-ol	1.29	1.67				

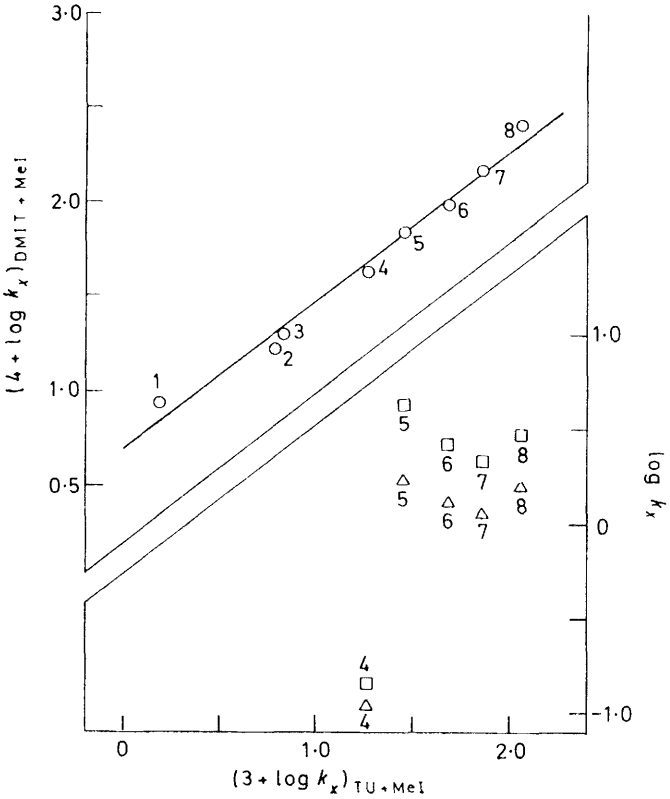


FIGURE 3 Empirical correlations for rate constants. ○, DMIT + MeI; □, TTT + MeI; △, DPTT + MeI, all plotted against TU + MeI. 1, Benzene; 2, propan-2-ol; 3, bromobenzene; 4, methanol; 5, cyclohexanone; 6, nitrobenzene; 7, acetonitrile; 8, propylene carbonate

Solvent Effects on the Rate Constant k_f.—The rate constants, k_f for the three reactions are summarized in Table 1 where k_x values calculated from the rate constant k_f and the molar volume of the solvent, Vs, expressed in l mol^{-1} by the equation $k_x = k_f V_s^{-1}$ are also given. There is a linear free-energy relationship between the two reactions both protic and aprotic solvents, *i.e.*, the reactions of *NN'*-dimethylimidazolidine-2-thione and of tetramethylthiourea with methyl iodide (Figure 3). The slope is 0.79. The result means that the reaction with the larger rate constant shows greater sensitivity to solvents.

In contrast, the reactions of the two triazole-3-thione derivatives TTT and DPTT with methyl iodide do not show any correlation with the tetramethylthiourea–methyl iodide reaction, although the TTT- and DPTT–methyl iodide reactions behave similarly. Thus, in two respects, *i.e.* the rate law in cyclohexanone and the solvent effects on rate constants, the TTT- and DPTT–methyl iodide reactions show abnormal characteristics. Further analysis is under way on this subject.

Pressure Effects and the Volume Change of Reaction.—Activation volumes and their pressure derivatives were determined by the least-squares method after expanding the rate constant into a power series of the pressure [equations (13)—(15)].

$$\ln k_p = \ln k_0 + Bp + Cp^2 \qquad (13)$$

$$\Delta V_0^{\ddagger} = (-1)\ BRT \qquad (14)$$

$$(\partial\ \Delta V^{\ddagger}/\partial p)_0 = (-2)\ CRT \qquad (15)$$

The apparent molal volumes of the solutes, ϕ_v, and the overall volume change in the reaction, $\Delta\phi_v$, are summarized in Table 3.

TABLE 2

Activation volumes and the pressure derivatives for the reaction of *NN'*-dimethylimidazolidine-2-thione with methyl iodide (30 °C)

Solvent	$\Delta V_0^{\ddagger}$/cm^3 mol^{-1}	$(\partial\,\Delta V^{\ddagger}/\partial p)_0$/cm^5 mol^{-1} kg^{-1}
Acetonitrile	−23.6	8.70 × 10^{-3}
Propylene carbonate	−16.9	4.57 × 10^{-3}

So far, molal volume measurements have been performed on the reactions of three thione derivatives with methyl iodide. Generally, observed trends are first, that molal volumes of thione derivatives and of ionic products are always larger in propylene carbonate than in acetonitrile, and second, that the reaction in acetonitrile gives larger negative volume changes than in propylene carbonate, on activation as well as on reaction.

TABLE 3

Apparent molal volumes ϕ_v and the volume change of the reaction at 30 °C

Solvent	ϕ_v/cm^3 mol^{-1}			$\Delta\phi_v$	ΔV_0‡
	DMIT	MeI	MDMIT$^+$I$^-$		
Acetonitrile	114.1	67.1	138.5	−42.7	−23.6
Propylene carbonate	119.5	65.1	162.3	−22.3	−16.9

Solvent properties which influence a molal volume of a solute might be reflected in other thermodynamic quantities of the solute. Thus, enthalpy measurements were performed in this connection.

Enthalpy Measurements.—Activation parameters have been calculated from the rate constant measured at four temperatures (40.0. 30.0, 20.0, and 0.0 °C). The enthalpy values are summarized in Table 4.

TABLE 4

Enthalpy change of solution, ΔH_s, and of reaction, ΔH_r, and activation parameters, $\Delta H^‡$ and $\Delta S^‡$

Solvent	ΔH_s/kcal mol^{-1}			ΔH_r/ kcal mol^{-1}	ΔH‡/ kcal mol^{-1}	ΔS‡/ cal K^{-1} mol^{-1}
	DMIT	MeI	MDMIT$^+$I$^-$			
Acetonitrile	4.69	0.869	3.07	−11.9	14.5	−25.1
Propylene carbonate	4.86	0.785	4.18	−9.44	14.2	−24.1

The enthalpy change accompanying the transfer of the reaction from acetonitrile to propylene carbonate calculated from the enthalpy change of reaction gives −9.4 − (−11.9) = 2.5 kcal mol^{-1}. The combination of the enthalpy change of solution using a thermodynamic cycle gives, for the same quantity, 1.0 kcal mol^{-1}. This discrepancy seems to be largely due to the difficulty in measuring the heat of reaction. The heats of solution are more reliable.

In contrast to our findings on the activation volumes, we did not observe any difference in practice in the activation enthalpies between the two solvents. As a probe searching for transition state properties, activation volumes are likely to be a more powerful tool, since they respond to the environment more sensitively.

A larger ΔH_s value in propylene carbonate than in acetonitrile is the usually observed trend for other quaternary ammonium halides.[15]

Transition State Index.—The volume change of the reaction, $\Delta\bar{V}^0$, as well as the activation volume, $\Delta V_0^‡$, have been regarded as consisting of two terms, *i.e.* the volume changes of the reactant molecules, $\Delta_1\bar{V}^0$ and $\Delta_1 V^‡$, and the volume changes of the surrounding solvent, $\Delta_2\bar{V}^\circ$ and $\Delta_2 V^‡$. In a previous paper,[12] ($\Delta_2 V^‡/\Delta_2\bar{V}^\circ$) values were suggested as charge development parameters along the reaction co-ordinate. The values $\Delta_1 V^‡$ −11.7 and $\Delta_1\bar{V}^\circ$ *ca.* 0 cm^3 mol^{-1} were also estimated on various grounds.[12] ($\Delta_2 V^‡/\Delta_2\bar{V}^\circ$) Values calculated on the same basis are given in Table 5.

A second transition state index, recently proposed is given by equation (16) [16,17] where E_a and ΔE stand for

$$n_T = E_a/(2\,E_a - \Delta E) \qquad (16)$$

the activation energy and energy change of the reaction. The index calculated from equation (16) by substituting corresponding enthalpy values for energies is also given in Table 5.

TABLE 5

Transition state indices for the reaction of *NN'*-dimethylimidazolidine-2-thione with methyl iodide (30 °C)

Solvents	($\Delta_2 V^‡/\Delta_2\bar{V}^0$)	n_T	α
Acetonitrile	0.28	0.35	0.43
Propylene carbonate	0.23	0.38	0.44

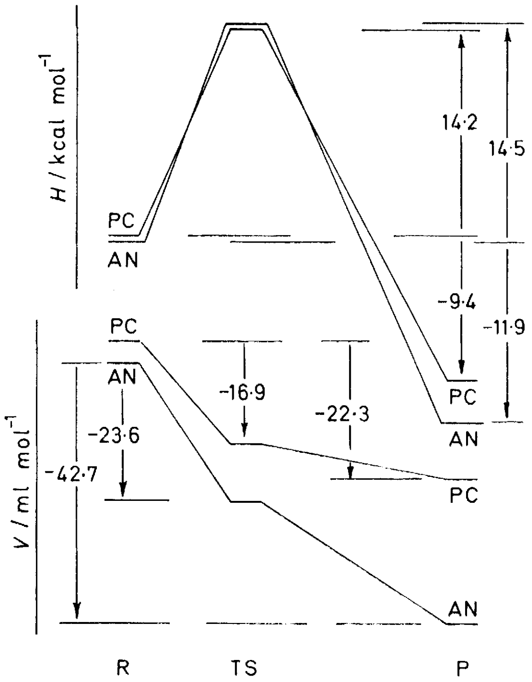


FIGURE 4 Volume and enthalpy *versus* reaction co-ordinate profile in acetonitrile (AN) and in propylene carbonate (PC)

Recently Albery extensively analysed nucleophilic substitution reactions by applying the Marcus equations and concluded that the Marcus expression holds for methyl transfer reactions.[18] According to his procedure one can obtain the transition state index, α, by two steps, when the values, $\Delta G^‡$ and ΔG°, are available, *i.e.* first solving equation (17) with respect to $\Delta G_0^‡$ and afterwards

$$\Delta G^‡ = \Delta G_0^‡ + (1/2)\,\Delta G^\circ + \Delta G^{\circ 2}/16\,\Delta G_0^‡ \qquad (17)$$

$$\alpha = (1/2)(1 + \Delta G^\circ/4\Delta G_0^‡) \qquad (18)$$

substituting the value into equation (18). The indices, calculated by substituting ΔH_r and $\Delta H^‡$ values into equations (17) and (18) instead of ΔG° and $\Delta G^‡$ values, are also included in Table 5. The three indices in

Table 5 are obtained with respect to the product pair of dissociated ions and not to the product ion-pair.

The first two indicate a transition-state position similar to that observed previously,[12] whereas the index α suggests a position more to the right. This is also the trend observed by a comparison of Abraham's Z parameter with the α index.[18]

One should notice here, however, that the pattern observed for the volume–reaction co-ordinate profile does not follow the Marcus pattern in the simplest way. In addition, activation enthalpy and enthalpy changes are detected as a result of compensation of the two large enthalpy changes accompanying bond-making and -breaking and solvation changes along the reaction co-ordinate during the process. In principle, however, in volume change analysis one can estimate volume changes accompanying the bond-making and -breaking, *i.e.* $\Delta_1V^\ddagger$ and $\Delta_1\bar{V}^\circ$, on some empirical basis. Thus the analysis allows us to derive the transition state index $(\Delta_2V^\ddagger/\Delta_2\bar{V}^\circ)$ which is free from the contribution.

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