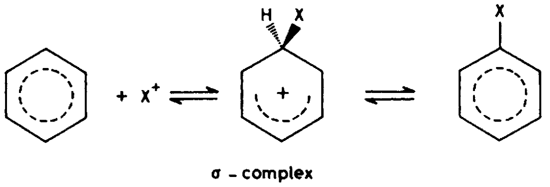
Protonation of 1,3,5-Tripyrrolidinobenzene in Basic Aqueous Solution. Thermodynamics and Kinetics of the Formation of a Stable σ-Complex

Siegmund Vogel, Wilhelm Knoche,* and Wolfgang W. Schoeller
Fakultät für Chemie, Universität Bielefeld, Postfach 8640, 4800 Bielefeld 1, West Germany

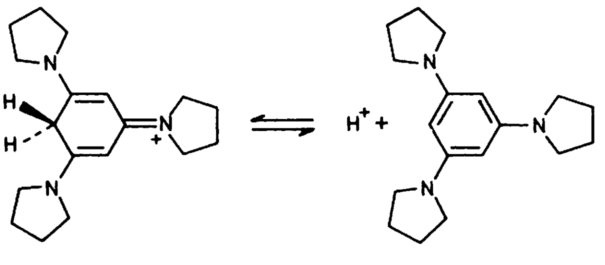
In aqueous solution, 1,3,5-tripyrrolidinobenzene is protonated at the aromatic ring, thus forming a σ-complex. The dissociation constant of this acid is $K_a = (2.50 \pm 0.25) \times 10^{-10}$ mol dm^{-3} in 9:1 v/v water–ethanol. The kinetics of protonation have been studied in basic solutions, where 1,3,5-tripyrrolidinobenzene reacts with water with the release of a hydroxide ion and free protons do not contribute to the reaction rate. The reaction is relatively slow and the rate constants are $k_1 = (0.62 \pm 0.04)$ s^{-1} for the protonation of TPB by water and $k_{-1} = (24\,300 \pm 400)$ cm^3 mol^{-1} s^{-1} for the reaction of the conjugate acid with hydroxide ions.

Electrophilic aromatic substitutions belong to the fundamental reactions in organic chemistry[1] and are of extreme importance. Their reaction mechanism is postulated as shown in Scheme 1. In the first step, an electrophile X^+ (NO_2^+, Hal^+, *etc.*) is added to the benzene ring and a σ-complex is formed, in which the cyclic π-conjugation is interrupted. In the second step, a proton leaves the σ-complex and the aromatic structure is regained. This two-step mechanism is now generally accepted. The mechanism was originally derived from chemical isotope studies,[2] which give indirect proof of the existence of a σ-complex as an intermediate (*i.e.* the reaction steps have not been observed separately). The existence of a σ-complex is also indicated by studies of protonation in superacidic media.[3]

Finally, in the case of 1,3,5-tripyrrolidinobenzene (TPB) Effenberger has found stable σ-complexes. Their salts can be isolated and their structures are well characterized.[4] In the most simple σ-complex the aromatic ring is protonated ($X^+ = H^+$). This complex was isolated as $TPBHClO_4$, and u.v. and n.m.r. spectra show clearly that protonation occurs at the aromatic ring and not at a nitrogen atom (see Scheme 2). Thermodynamics and kinetics of the reaction are reported upon in this contribution.



Scheme 1. Electrophilic aromatic substitution



Scheme 2.

Experimental

A Cary 219 instrument was used for the spectrophotometric measurements. The kinetic data were obtained with pressure-jump equipment using optical detection and the data handling system from Dialog. Relaxation times were evaluated in the usual way and the relaxation amplitudes were obtained by the following calculation. The equipment permits the measurement of the intensity of light, I, behind the sample cell and its change δI caused by the chemical relaxation effect. It is assumed that only a single species absorbs light at the wavelength applied. Then the absorbance A is given by equation (1) where ε is the

$$A = \varepsilon c l \qquad (1)$$

absorbance coefficient, c the concentration of the absorbing species, and l the optical length of the absorbing solution. In a relaxation experiment, the pressure drops from a value P_x to $P = 1$ atm at time $t = 0$. As a result, the concentration changes from $\bar{c} + x^\circ$ at $t = 0$ to $\bar{c}$ at $t \gg \tau$, when the reaction has reached equilibrium. The absorbance changes correspondingly from $\bar{A} + \delta A$ to $\bar{A}$. Lambert–Beer's law permits the calculation of the change in light intensity caused by the change is absorbance [equation (2) where I is the light intensity at

$$\delta I = I\,(10^{-\delta A} - 1) \qquad (2)$$

equilibrium]. Spectrophotometric measurements are performed in the range A *ca.* 1, and in relaxation experiments there is always $\delta A \ll 1$. Therefore equation (2) can be approximated by (3). I and δI are measured in an experiment, ε is known from

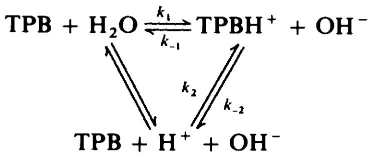
$$\delta I = -2.3\, I \varepsilon l x^\circ \qquad (3)$$

equilibrium measurements, the length of the optical cell in the equipment used is $l = 22$ mm, and therefore the change in concentration of the absorbing species, x°, can be calculated.

pH Was measured using a standard glass electrode calibrated at 25 °C with 0.01N-borax (pH 9.18), a mixture of 0.025N-borax (50 ml) and 0.1N-NaOH (12.2 ml) (pH 10.10), and a mixture of 0.025N-borax (50 ml) and 0.1N-NaOH (19.1 ml) (pH 10.60).[5] In all experiments the solutions were thermostatted within ±0.1 K. The TPB was prepared as salt $TPBHClO_4$ according to reference 6 and kindly provided by Professor F. Effenberger, University of Stuttgart. All other reagents were analytical grade. Since TPB is only slightly soluble in aqueous solution, at first a 5×10^{-4} molar solution was prepared in ethanol or acetonitrile. Water and organic solvent was added to reach the final concentration of 5×10^{-5}M. Nevertheless, the solubility is greatly restricted in basic solutions, where the unprotonated TPB prevails. The pH was adjusted by adding 10^{-3}M-NaOH. CO_2 Had to be removed by carefully degassing the solvent, and the measurements had to be performed under nitrogen gas, since unbuffered solutions were used at pH 9. Solutions were freshly prepared before use, as TPB decomposes slowly.

Results and Discussion

Relaxation Times and Amplitudes.—In basic solution, TPB may react either with H^+ or with H_2O under release of OH^-. Therefore, the reaction scheme may be written as (4). The



(4)

dissociation of water is too fast to be observed using the pressure-jump technique, and therefore a single relaxation effect is expected for equation (4). The relaxation time is obtained from the rate law (5). Activity coefficients can be neglected at the

$$-\mathrm{d}[\mathrm{TPB}]/\mathrm{d}t = k_1[\mathrm{TPB}] - k_{-1}[\mathrm{TPBH}^+][\mathrm{OH}^-] + k_2[\mathrm{TPB}][\mathrm{H}^+] - k_{-2}[\mathrm{TPBH}^+] \qquad (5)$$

low ionic strength used (*I ca.* 10^{-4} mol dm^{-3}). The water activity does not change during the progress of the reaction and is included in k_1. In order to integrate equation (5), the actual concentrations $[A]$ are replaced by the equilibrium concentrations $[\bar{A}]$ and the deviations from equilibrium x_A [equation (6)]. All equilibrium terms cancel each other out in equation (5).

$$[A] = [\bar{A}] + x_A \qquad (6)$$

In chemical relaxation experiments, we always have relationship (6a) and therefore terms may be neglected in which x is squared.

$$x_A \ll [A] \qquad (6a)$$

Finally, all x_A have to be replaced by x_{TPB}. This is achieved by (i) the mass balance (7) of TPB; (ii) the condition (8) of electroneutrality; (iii) the fact that water dissociation is very fast and therefore equilibrium (9) is established for this reaction.

$$x_{\mathrm{TPB}} + x_{\mathrm{TPBH}} = 0 \qquad (7)$$

$$x_{\mathrm{TPBH}} + x_{\mathrm{H}} = x_{\mathrm{OH}} \qquad (8)$$

$$([\overline{\mathrm{OH}^-}] + x_{\mathrm{OH}})([\overline{\mathrm{H}^+}] + x_{\mathrm{H}}) = [\overline{\mathrm{OH}^-}][\overline{\mathrm{H}^+}] \qquad (9)$$

Inserting equations (6)—(9) into (5) yields (10) with the relationship (11). For pH > 8, where $[\mathrm{H}^+] \ll [\mathrm{OH}^-]$, this may be reduced to (12). Equations (12) shows that for $[\mathrm{OH}^-] > [\mathrm{TPBH}^+]$ the relaxation rate $1/\tau$ increases with pH if TPB reacts predominantly with water, and that the relaxation rate decreases with an increase of pH if TPB reacts predominantly with hydrogen ions.

$$\mathrm{d}x_{\mathrm{TPB}}/\mathrm{d}t = -\frac{1}{\tau}x_{\mathrm{TPB}} \qquad (10)$$

$$\frac{1}{\tau} = k_1 + k_{-1}\left([\mathrm{OH}^-] + \frac{[\mathrm{TPBH}^+][\mathrm{OH}^-]}{[\mathrm{H}^+] + [\mathrm{OH}^-]}\right) + k_2\left(\mathrm{H}^+] + \frac{[\mathrm{TPB}][\mathrm{H}^+]}{[\mathrm{H}^+] + [\mathrm{OH}^-]}\right) + k_{-2} \qquad (11)$$

$$\frac{1}{\tau} = k_1 + k_{-1}([\mathrm{OH}^-] + [\mathrm{TPBH}^+]) + k_2([\mathrm{H}^+] + [\mathrm{TPB}][\mathrm{H}^+][\mathrm{OH}^-]^{-1}) + k_{-2} \qquad (12)$$

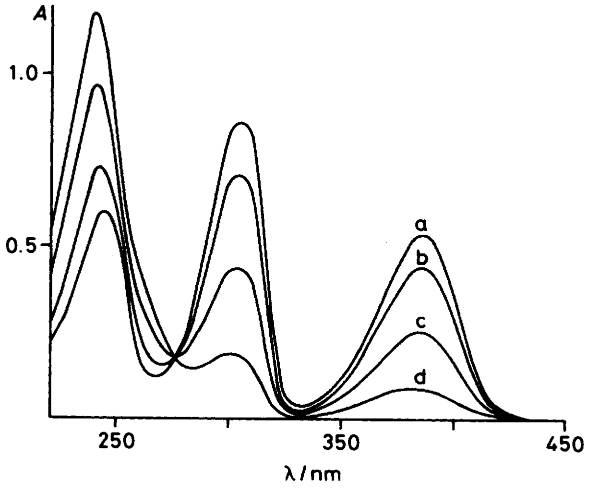


Figure 1. Spectra of 2.60 × 10^{-5}M solutions of $\mathrm{TPBHClO_4}$ in 9:1 v/v water–ethanol at 25 °C: a, pH 7.30; b, pH 9.05; c, pH 9.65; d, pH 10.30

The rate constants are related to the equilibrium constants K_a and K_b by equations (13) and (14). K_aK_b Is the ion product of water.

$$\frac{k_1}{k_{-1}} = \frac{[\mathrm{TPBH}^+][\mathrm{OH}^-]}{[\mathrm{TPB}]} = K_b \qquad (13)$$

$$\frac{k_{-2}}{k_2} = \frac{[\mathrm{TPB}][\mathrm{H}^+]}{[\mathrm{TPBH}^+]} = K_a \qquad (14)$$

The equilibrium of the reaction will be determined from the amplitude of the relaxation effect. K_b Is altered by a pressure change according to van't Hoff's equation (15). This leads to

$$\left(\frac{\partial \ln K_b}{\partial P}\right)_T = -\frac{\Delta V^\circ}{RT} \qquad (15)$$

$$\frac{x^\circ_{\mathrm{TPB}}}{[\mathrm{TPB}]} + \frac{x^\circ_{\mathrm{TPBH}}}{[\mathrm{TPBH}^+]} + \frac{x^\circ_{\mathrm{OH}}}{[\mathrm{OH}^-]} = -\frac{\Delta V^\circ \delta P}{RT} \qquad (16)$$

(16) where x°_i is the change in concentration of species i caused by a pressure change δP. The chemical relaxation is observed at 387 nm where only TPBH^+ absorbs light. x°_{TPBH} Is obtained by combining equation (16) with (7)—(9) where Γ is given by equation (18).

$$x^\circ_{\mathrm{TPBH}} = \Delta V^\circ \cdot \frac{\delta P}{RT} \cdot \Gamma \qquad (17)$$

$$\Gamma = \left(\frac{1}{[\mathrm{TPB}]} + \frac{1}{[\mathrm{TPBH}^+]} + \frac{1}{[\mathrm{OH}^-]}\right)^{-1} \qquad (18)$$

Static Measurements.—Figure 1 shows the spectra of solutions of TPB at pH 7.3—10.3, clearly indicating the protonating step occurring in this region. The deprotonated species shows absorption maxima at 244 (ε 55 700) and 317 nm (2 100), the protonated species at 245 (ε 22 100), 306 (33 300), and 387 nm (20 800). The numbers in brackets are the values of the absorbance coefficient in cm^{-1} dm^3 mol^{-1}. The absorbance at 387 nm has been used to determine the protonation equilibrium.

Table 1. pK_a Values obtained from different measurements. Solvent: 9:1 v/v water–ethanol

Measurements	20 °C	25 °C
Spectrophotometric	9.71 ± 0.03	9.62 ± 0.03
Relaxation time	9.66 ± 0.05	9.60 ± 0.05
Relaxation amplitude	9.72 ± 0.06	9.62 ± 0.06

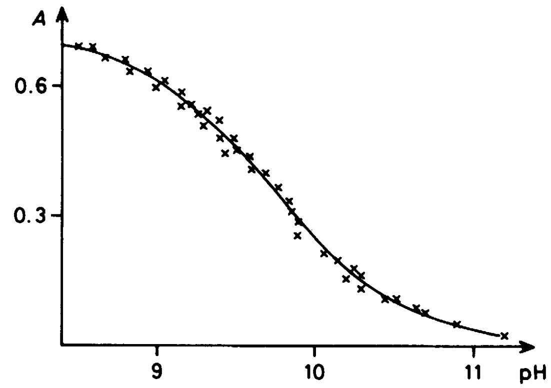


Figure 2. Absorbance at 287 nm *versus* pH for 3.37 × 10^{-5}M solutions of $TPBHClO_4$ in 4:1 v/v water–ethanol at 25 °C. The curve is calculated for a single protonation with pK_a 9.62

As an example, Figure 2 shows the absorbance as a function of pH for a solvent with the composition 4:1 v/v water–ethanol. The data fit to a single protonation step (calculated curve in Figure 2) and the pK_a can be evaluated to an accuracy of ±0.03. The deviations in Figure 2 are caused mainly by the slow decomposition of TPB.

Nearly all measurements are performed in unbuffered solutions with ionic strength *I ca.* 10^{-4} mol dm^{-3} and activity coefficients are neglected in calculating K_a. The composition of the solvent has been changed in order to check the influence of the inorganic compound of the solvent on the equilibrium. The results are summarized in Table 2.

In 9:1 water–ethanol, K_b has been determined in the temperature range 5—40 °C. The data are described by $d\ln K_b/dT^{-1} = (950 \pm 50)$ K.

Kinetic Measurements.—The progress of the reaction is followed at λ = 387 nm. At constant TPB concentration, the rate increases with increasing pH. Therefore, it is concluded that TPB reacts with H_2O rather than with H^+, and equation (12) is reduced to (19). $[TPBH^+]$ is determined spectrophotometrically from the absorbance of the solution at λ 387 nm.

$$\frac{1}{\tau} = k_1 + k_{-1}([OH^-] + [TPBH^+]) \qquad (19)$$

The pH of the solution is measured and $[OH^-]$ is determined as $K_w 10^{pH}$. We rewrite equation (19) and obtain (20).

$$\frac{1}{\tau} = k_1 + k_{-1}K_w(10^{pH} + [TPBH^+]K_w^{-1}) \qquad (20)$$

In Figure 3 some results are plotted according to equation (10), yielding k_1 as intercept and $k_{-1}K_w$ as slope. k_{-1} Can be calculated, since K_w is known from the literature.[7] The equilibrium constant K_a can be determined from the kinetic measurements according to equation (21).

$$K_a = k_{-1}K_w/k_1 \qquad (21)$$

The measurement of pH values becomes difficult when the organic compound exceeds 20 vol %. For these solvents $[OH^-]$ is determined as the concentration of NaOH added to the solution and the relaxation times are evaluated using equation (19).

Table 2. Equilibrium and rate constant at 25 °C

Solvent* (vol%)	pK_b	k_1/s^{-1}	10^4k_{-1}/mol dm^{-3} s^{-1}
10% ethanol	4.60 ± 0.04	0.63 ± 0.05	2.5 ± 0.1
20% ethanol	4.69 ± 0.05	0.73 ± 0.08	3.5 ± 0.2
30% ethanol	4.80 ± 0.06	0.85 ± 0.08	5.4 ± 0.2
40% ethanol	4.99 ± 0.08	1.00 ± 0.12	9.8 ± 0.4
5% acetonitrile	4.46 ± 0.06	0.69 ± 0.05	2.0 ± 0.2
20% dioxane	4.66 ± 0.06	0.74 ± 0.06	3.3 ± 0.3

* Water is the second component of the solvent.

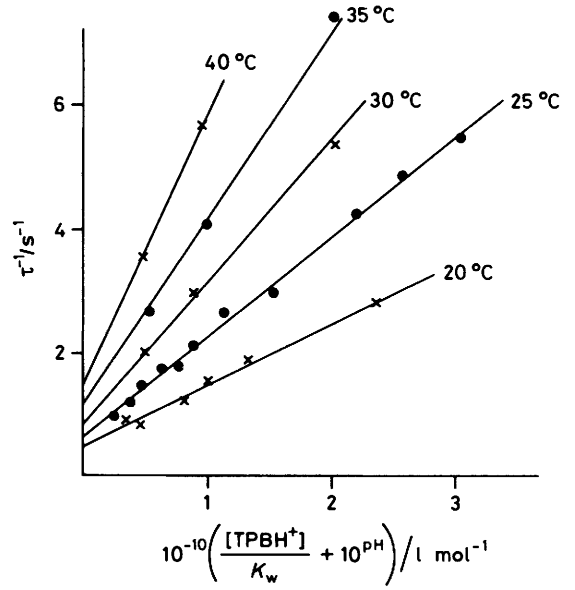


Figure 3. Relaxation rate plotted according to equation (20) for 1.66 × 10^{-5}M solutions of $TPBHClO_4$ in 9:1 v/v water–ethanol

K_a May also be determined from the relaxation amplitude. The amplitude reaches its maximum value when the transfer function Γ is at its maximum level and Γ [equation (18)] may be expressed as a function of K_a and the measurable quantities $[TPBH^+]$ and pH. Figure 4 shows a plot of the amplitude against pH at constant concentration of TPB, together with a curve which is proportional to Γ as calculated according to equation (18) with K_a = 9.62. The three independently obtained values of K_a are summarized for two temperatures in Table 1. The conformity of these values clearly indicates that the relaxation effect is due to the protonation reaction of TPB.

The values for k_1 and k_{-1} are summarized for different solvent composition in Table 2. For 9:1 v/v water–ethanol, they have been determined between 20 and 40 °C. The results are described by $d\ln k_1/dT^{-1} = (5.8 \pm 0.3) \times 10^3$ K and $d\ln k_{-1}/dT^{-1} = (5.4 \pm 0.4) \times 10^3$ K. A further check of the reaction scheme is provided by the dependence of the reaction rate on ionic strength. We assume that the activated complex is uncharged and that activity coefficients are unity for uncharged species, and we obtain equation (22) instead of (19). $f_\pm$ Is the

$$\frac{1}{\tau} = k_1 + k_{-1}f_\pm^2([OH^-] + [TPBH^+]) \qquad (22)$$

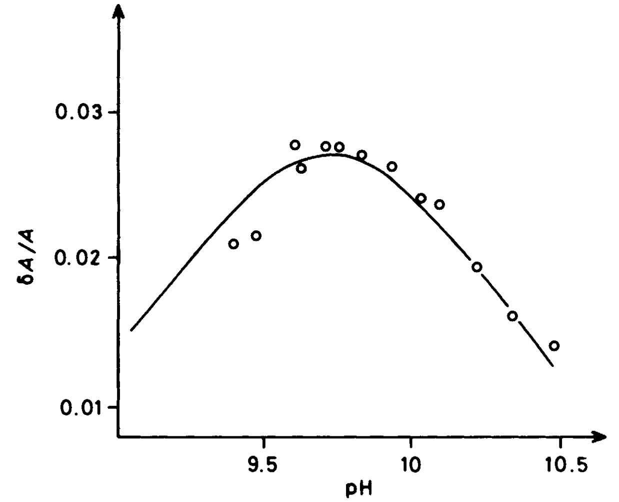


Figure 4. Amplitudes of the relaxation effect *versus* pH for 1.66×10^{-5}M solutions of $TPBHClO_4$ in 9:1 v/v water–ethanol. The curve is calculated according to equation (17) with pK_a 9.62

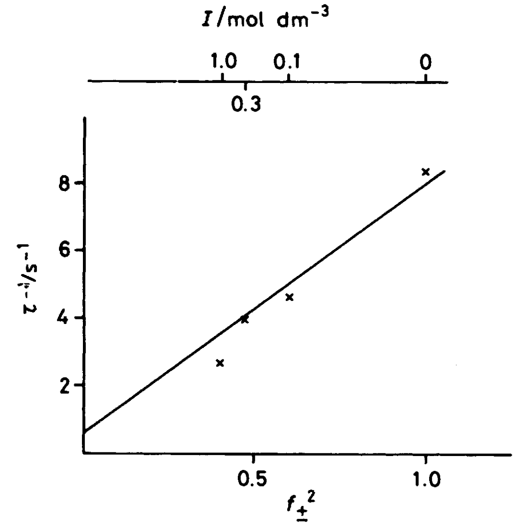


Figure 5. Relaxation rates *versus* $f_{\pm}^2$ for 1.66×10^{-5}M solutions of $TPBHClO_4$ in 9:1 v/v water–ethanol. The ionic strength I of the solution is altered by adding NaCl. $f_{\pm}$ Is taken from ref. 5 and the straight line is calculated by equation (22): T 25 °C; $[OH^-]$ 3×10^{-4}M

mean activity coefficient. The ionic strength is changed by adding different amounts of NaCl to the solutions. The expected dependence is observed as shown in Figure 5 for constant concentrations of OH^- and $TPBH^+$.

Generally, proton transfer occurs at a diffusion-controlled rate.[8] However, this is not true for carbon acids, where rate constants *ca.* 10^4 dm^3 mol^{-1} s^{-1} have been reported.[9] A similar value is found in this investigation for k_{-1}. This shows clearly that TPB acts as a carbon base, which protonates at the aromatic ring and not at a nitrogen atom.

The reaction of TPB with H^+ in reaction (4) is not observed. This permits a calculation of an upper limit for the rate constants of this reaction (k_2,k_{-2}) by applying equation (12) to the slowest relaxation rate observed. This yields $k_2 < 5 \times 10^7$ dm^3 mol^{-1} s^{-1}, $k_{-2} < 8 \times 10^{-3}$ s^{-1} at 25 °C, assuming that the maximum error of τ is 10%.

Table 2 shows that the equilibrium and rate constants depend slightly upon the amount of organic solvent. It may be concluded that the organic solvent is not involved in the reaction, since similar results are obtained for the addition of such varying solvents as ethanol and dioxane. The water activity varies between 0.99 and 0.90 for the solvent mixtures used (see Table 2). Therefore, it may be expected that k_1 depends little upon the solvent composition. However, a large change in k_1 is observed, which cannot be explained. k_{-1} Increases when organic solvent is added. Again, so far an explanation of this behaviour cannot be given, since dioxane and ethanol both influence k_{-1} in a similar way.

Table 3. Results for reaction of $TPB + H_2O \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} TPBH^+ + OH^-$. Water activity = 1, T = 25 °C

$$\begin{aligned}
K_b &= 10^{-4.59 \pm 0.05} \text{ mol dm}^{-3} \\
\Delta H^{\circ} &= (8 \pm 2) \text{ kJ mol}^{-1} \\
\Delta S^{\circ} &= (-60 \pm 8) \text{ J mol}^{-1} \text{ K}^{-1} \\
\Delta V^{\circ} &= (-12 \pm 1) \text{ cm}^3 \text{ mol}^{-1} \\
k_1 &= (0.62 \pm 0.04) \text{ s}^{-1} \\
k_{-1} &= (24.3 \pm 0.3) \times 10^3 \text{ dm}^3 \text{ s}^{-1} \text{ mol}^{-1} \\
\Delta H_1{}^{\pm} &= (46 \pm 3) \text{ kJ mol}^{-1} \\
\Delta S_1{}^{\pm} &= -(96 \pm 10) \text{ J mol}^{-1} \text{ K}^{-1} \\
\Delta H_{-1}{}^{\pm} &= (42 \pm 3) \text{ kJ mol}^{-1} \\
\Delta S_{-1}{}^{\pm} &= -(20 \pm 10) \text{ J mol}^{-1} \text{ K}^{-1}
\end{aligned}$$

From the temperature dependence of K_b the reaction enthalpy $\Delta H°$ is 8 kJ mol^{-1} and the reaction entropy $\Delta S°$ is -60 J mol^{-1} K^{-1}. The reaction volume $\Delta V°$ -12 cm^3 mol^{-1} is obtained according to equation (17) from the amplitude of the relaxation effect. The temperature dependence of k_{-1} yields activation enthalpies and entropies: $\Delta H_1{}^{\ddagger}$ 46 kJ mol^{-1}, $\Delta S_1{}^{\ddagger}$ -96 J mol^{-1} K^{-1}, $\Delta H_{-1}{}^{\ddagger}$ 42 kJ mol^{-1}, $\Delta S_{-1}{}^{\ddagger}$ -20 J mol^{-1} K^{-1}. Finally, the dependence of the reaction rate on ionic strength clearly indicates that the activated complex is uncharged. All results are summarized in Table 3.

In basic aqueous solution the results may be summarized as follows. 1.3.5-Tripyrrolidinobenzene is protonated according to equilibrium (23) and free protons do not contribute to the

$$TPB + H_2O \rightleftharpoons TPBH^+ + OH^- \qquad (23)$$

reaction rate. So far a more refined discussion of the reaction mechanism and of the structure of the activated complex is not possible. In order to obtain more information, further measurements are being performed to study the influence of substituents on equilibrium and rate of the reaction.

Acknowledgements

We thank Professor F. Effenberger for providing TPB, and the Fonds der Chemischen Industrie for support.

References

1 C. H. Bamford and C. F. H. Tipper, 'Comprehensive Chemical Kinetics,' vol. 13, 'Reactions of Aromatic Compounds,' Elsevier, Amsterdam, 1972.
2 L. Melander, 'Isotope Effects on Reaction Rates,' Ronald Press, New York, 1960.
3 H. Köhler and A. Scheibe, *Z. Anorg. Allg. Chem.*, 1956, **285**, 221; T. Yamaoka, H. Hosoya, and S. Nagakura, *Tetrahedron*, 1970, **26**, 4125; G. A. Olah, R. H. Schlosberg, R. D. Porter, Y. K. Mo, D. P. Kelley, and G. D. Mateescu, *J. Am. Chem. Soc.*, 1972, **94**, 2034.
4 F. Effenberger, K. E. Mack, K. Nagel, and R. Niess, *Chem. Ber.*, 1977, **110**, 165.
5 R. A. Robinson and R. H. Stokes, 'Electrolyte Solutions,' Butterworth, London, 1970, p. 544.
6 F. Effenberger and R. Niess, *Angew. Chem., Int. Ed. Engl.*, 1967, **6**, 1061.
7 B. Gutbezahl and E. Grunwald, *J. Am. Chem. Soc.*, 1953, **75**, 565.
8 M. Eigen, *Angew. Chem.*, 1963, **75**, 489; *Angew. Chem., Int. Ed. Engl.*, 1964, **3**, 1.
9 R. P. Bell, 'The Proton in Chemistry,' Chapman and Hall, London, 1973.

Received 5*th September* 1985; *Paper* 5/1516