

TAUTOMERISM OF 3-THIANAPHTHENONE: SPECTRA, EQUILIBRIA AND KINETICS

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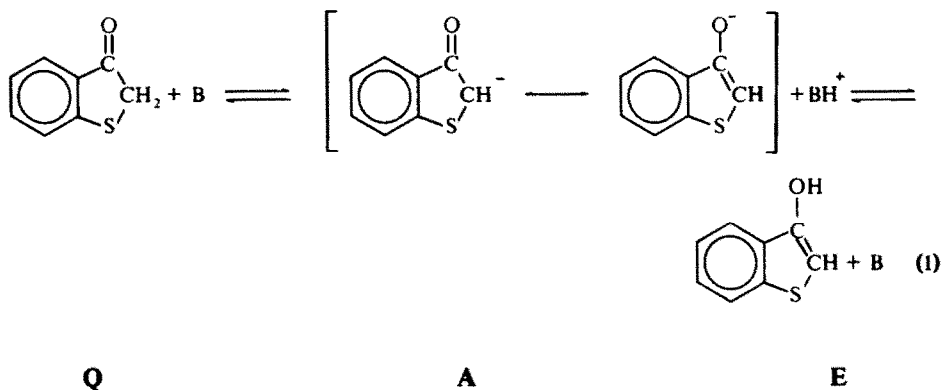
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Abstract—UV absorption spectra of a tautomeric system (1) of 3-thianaphthenone (3-TN) were measured and the spectra of individual chemical species involved were determined (Figs 1 and 2). On acidifying an alkaline solution of 3-TN practically pure enol was formed in the first stage of reaction. The tautomeric equilibrium constant [Enol]/[Ketone] was found to be 0.18 in H₂O, 2.02 in ethanol. In aqueous solution, the p*K* of the enol was 8.97, and the p*K* of the ketone was 9.71. The rate constants of proton transfer in the ketone-anion system were $7.9 \times 10^{-4} \text{ sec}^{-1}$ for dissociation, and $4.05 \times 10^6 \text{ M}^{-1}\text{sec}^{-1}$ for recombination. Corresponding rate constants for the enol-anion system were of the order of 10^8 – $10^9 \text{ M}^{-1}\text{sec}^{-1}$ and 0.1 – 1 sec^{-1} , respectively. Freshly sublimed 3-TN is in the ketone form, and tautomerizes extremely slowly in crystals or in aprotic solvents. The results explain apparent contradictions between NMR spectra of 3-TN described by different authors, and the anomalous polarographic behaviour of the compound.

INTRODUCTION

3-THIANAPHTHENONE (3-TN) (2,3-dihydrothianaphthen-3-on; benzol[*b*]thiophen-3(2*H*)-on) may exist in two tautomeric forms, as a ketone, **Q**, or an enol, **E** (thioindoxyl; benzol[*b*]thiophene-3-ol). The tautomeric equilibrium (1)



is usually described as being shifted to the left, as indicated by a bromometric titration of an aqueous solution of 3-TN,¹ by IR^{2,3} and UV spectra.⁴ Nevertheless, Buu-Hoï *et al.*⁵ found only the enol form by measuring the NMR spectra of 3-TN in CDCl₃ or in CF₃COOH. Pure tautomers could not be isolated, but two kinds of reactivity of 3-TN were described.⁶⁻⁸

Zuman *et al.*⁹ studied the equilibrium (1) by the polarographic method, and determined the apparent ionization constant of 3-TN, and found that the reactions $\text{Q} + \text{B} = \text{A} + \text{BH}^+$ are very slow from the point of view of polarography, i.e. their relaxation

times are $\tau \gg 1$ sec. Inspired by Zuman, we tried to determine the protolytic equilibria and the rates of tautomerization of 3-TN by means of UV spectrophotometry.

Spectra of the tautomeric system and of the individual tautomers

The UV absorption spectrum of 3-TN in aqueous (containing 5% of ethanol) solution in tautomeric equilibrium (Fig. 1, curve *a*) consists of four distinct bands with the following λ_{\max} and $\log_{10} \epsilon_{\max}$ (the last value given in parentheses): I—370 nm (3.40); II—303 nm (2.80) and 295 nm (2.75); III—267 nm (3.67) and 260 nm (3.72); IV—235 nm (4.3).

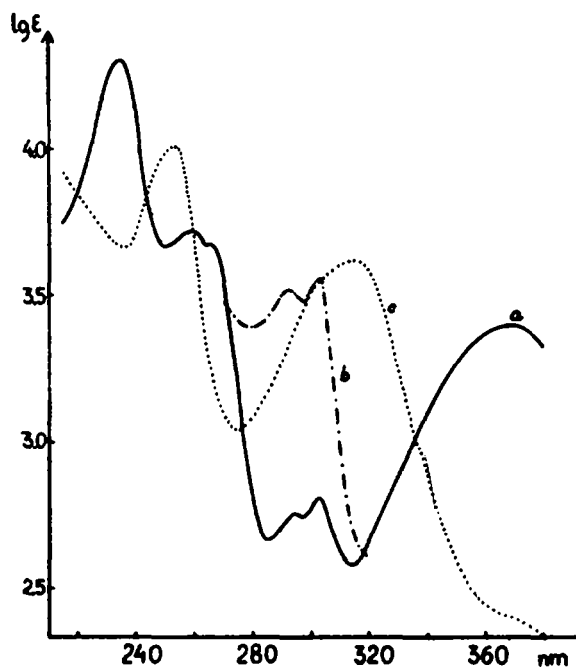


FIG. 1 UV absorption spectra of 3-TN: *a*—equilibrated tautomeric mixture in H_2O (+5% ethanol); *b*—enol E, in the same solvent from kinetic extrapolation to $t=0$; *c*—anion A, in borate buffer, $pH > 12$.

Based on IR and NMR spectra, we found that freshly sublimed 3-TN is practically pure ketone Q. Therefore, the absorption spectrum of such a sample freshly dissolved in hexane (fig. 2, curve *a*) represents the spectrum of Q, because of the time required to establish a tautomeric equilibrium in this aprotic and nonpolar solvent. In this spectrum the band II is absent; in ethanol solution the same band is intense (Fig. 2, curve *b*), while the band I is weakened. This enabled us to ascribe the absorption band I to the ketone Q, the band II—to the enol E.

The absorption spectrum of the anion A (Fig. 1, curve *c*) differs considerably from that of a neutral molecule, and reveals a shoulder at $\lambda \approx 370$ nm (which is not due to traces of Q), and two maxima, at λ 315 nm (3.62) and 254 nm (4.00).

When acid is added to an alkaline solution of 3-TN, the enol form E is formed in practically a 100% yield and subsequently decays. By kinetic extrapolation of a logarithmic plot of absorbancy at a given λ vs time to $t = 0$ (analogously to previous work¹⁰), we determined part of the spectrum of the enol, as shown in Fig. 1, curve *b*, in the region of 300 nm where the absorption of Q is low. This absorption band of E is very similar to that appearing in a spectrum of 3-methoxythianaphthene.^{4,9}

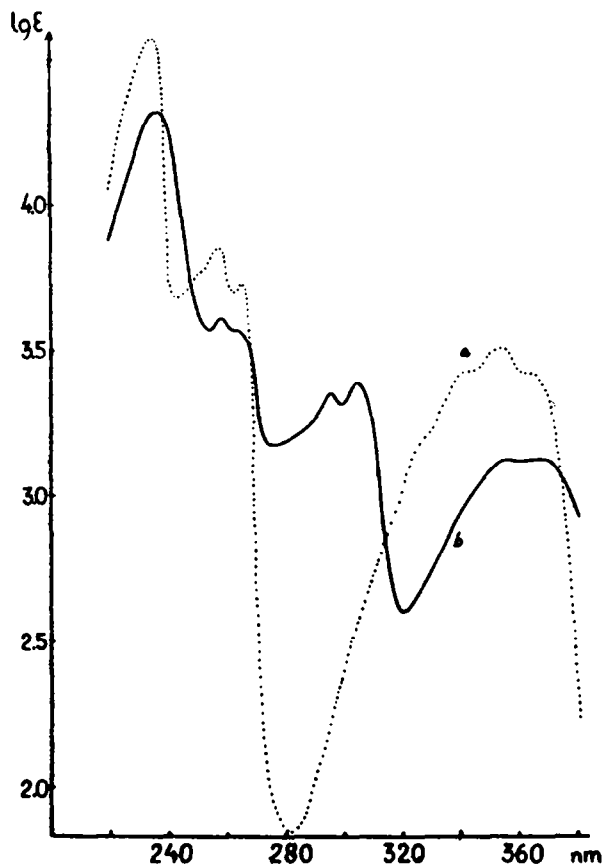


FIG. 2 UV absorption spectra of 3-TN: *a*—freshly sublimed 3-TN, solution in *n*-hexane; *b*—ethanol solution, in a tautomeric equilibrium.

Absorption coefficients of ketone Q in the same spectral region were obtained by successive approximations, assuming the shape of the spectrum of Q in aqueous solution to be analogous to that in hexane (Fig. 2, curve *a*).

The NMR spectrum of 3-TN reveals two resonance absorptions: a single signal at $\delta = 3.75$ ppm from internal TMS standard, and a complex pattern at $\delta \approx 7.4$ ppm. The latter integrated signal is twice as large as the former, which corresponds to 4 nonequivalent protons in the aromatic ring, and 2 protons in a CH_2 group of Q. No signal was found in the region characteristic of OH protons.

The IR spectrum of sublimed 3-TN in CHCl_3 solution shows the strongest absorption at 1695 cm^{-1} ($\text{C}=\text{O}$ stretching vibration), and strong or moderately strong bands at 1600, 1580, 1455, 1393, 1310, 1290 and 1075 cm^{-1} . No absorption was found in the region $1800\text{--}4000\text{ cm}^{-1}$.

Tautomeric equilibrium

The tautomeric equilibrium constant of 3-TN, $K_T = [\text{E}]/[\text{Q}]$, was calculated from the spectrophotometric determination of $[\text{E}]$ and $[\text{Q}]$ at equilibrium, using the absorption coefficients found for both forms as described above. $K_T = 0.18 \pm 0.01$ at 20° in aqueous 5% ethanol, as a mean for three solutions measured at four different wavelength (Table 1).

TABLE 1. DETERMINATION OF TAUTOMERIC EQUILIBRIUM CONSTANT K_T FOR SEVERAL AQUEOUS SOLUTIONS OF 3-TN (+5% ETHANOL)

λ nm	ϵ_E $\text{M}^{-1}\text{ cm}^{-1}$	ϵ_Q	K_T		
295	3200	90	0.177	0.191	0.184
298	3000	110	0.178	0.195	0.183
300	3140	130	0.172	0.192	0.181
303	3560	140	0.170	0.190	0.175

mean: $K_T = 0.18 \pm 0.01$

The tautomeric equilibrium constant determined in the same manner for pure ethanol solution is $K_T(\text{EtOH}) = 2.02 \pm 0.12$.

Protolytic equilibria

The optical density (absorbance), $D(\lambda)$, of the system is given by

$$D(\lambda) = \epsilon_Q(\lambda)[\text{Q}] + \epsilon_E(\lambda)[\text{E}] + \epsilon_A(\lambda)[\text{A}]. \quad (2)$$

For equilibrium, we substitute:

$$[\text{A}]_\infty = K_Q[\text{Q}]_\infty/[\text{H}^+] \quad \text{and} \quad [\text{E}]_\infty = K_T[\text{Q}]_\infty. \quad (3)$$

Under conditions of the experiment ($\text{pH} \leq \text{p}K_Q$) we can neglect the last term in Eq. (2) for $\lambda \approx 370\text{ nm}$ (see Fig. 1). Formulating an equation for a ratio of optical densities at two different wavelengths, $\lambda_1 \approx 300\text{ nm}$ with high ϵ_E and ϵ_A , and for $\lambda_2 \approx 370\text{ nm}$ with high ϵ_Q and low ϵ_A , we obtain a straight line equation:

$$D(\lambda_1)/D(\lambda_2) \{ \epsilon_Q(\lambda_2) + \epsilon_E(\lambda_2)K_T \} \simeq \epsilon_Q(\lambda_1) + \epsilon_E(\lambda_1) + \epsilon_A(\lambda_1)K_Q/[\text{H}^+] \quad (4)$$

From the measurements of the left side term of Eq. (4) at different pH's we calculated by the least square method the value of the ionization constant of the ketone:

$$K_Q = 1.95 (\pm 0.09) \times 10^{-10}\text{ M.}$$

and, consequently, the ionization constant of the enol:

$$K_E = K_Q/K_T = 1.07 (\pm 0.05) \times 10^{-9}\text{ M.}$$

Tautomerization kinetics

Tautomeric change, as a prototropic reaction, proceeds usually as a sequence of proton transfer steps. In sufficiently dilute aqueous solutions* the overall kinetics may be represented by a general scheme:



where $k_j = \sum_i k_{ij}[B_i]$ for $j = 2$ and 4 (6a)

and $k_j = \sum_i k_{ij}[\text{HB}_i] = [\text{H}^+] \sum_i k_{ij}K_{\text{HB}_i}[B_i]$ for $j = 1$ and 3 (6b)

where B_i and HB_i stand for i -th base and its conjugate acid.

To simplify the analysis of the kinetics, borate concentrations were kept low, so that participation of a borate ion and of boric acid in proton transfer reactions at $\text{pH} \approx 2.3$ could be neglected, and the only significant acid and base were H^+ or H_2O , respectively:

$$k'_j = k_j[\text{H}^+] \quad \text{for } j = 1 \text{ and } 3. \quad (7)$$

If excess of H^+ ions assures a pseudo-first order of reaction, the system of differential kinetic equations describing the reaction (5) has a general solution,¹² involving two relaxation times of reaction.

Two relaxation times correspond to two stages of the reaction (5). However, the kinetic measurements revealed only one relaxation time (Fig. 3); this indicates that an equilibrium in the anion-enol system is established, within a few seconds, prior to the first measurement.

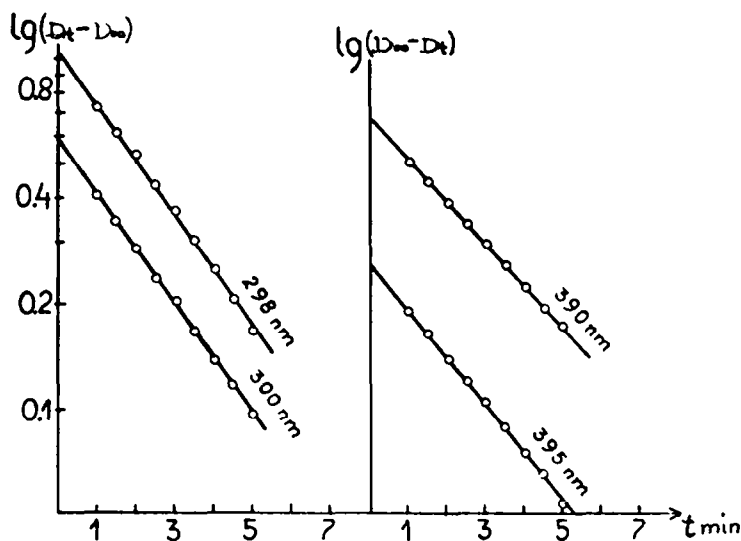


FIG. 3 Examples of relaxational plots of changes in optical density, $\lg_{10} |D(\lambda, t) - D(\lambda, \infty)|$ against time, for a region of absorption of enol, $\lambda \approx 300$ nm, and that of ketone, $\lambda \approx 390$ nm.

* In concentrated solutions the investigated compound itself catalyses the reaction, and the kinetics become much more complex.¹¹

Assuming this equilibrium ($K_E = k_2/k_1$) to be established, and substituting $d[Q]/dt = -d([E] + [A])/dt$ into a relaxational equation,¹³ we obtain a single relaxation time τ given by

$$1/\tau = (k_2k_3' + k_2k_4 + k_1'k_4)/(k_1' + k_2). \quad (8)$$

As $k_1' \gg k_2$ (at pH 2.3 $k_1' = k_1[H^+] = k_2[H^+]/K_E > 10^6 k_2$), we neglect k_2 in the denominator of (8) and obtain:

$$1/\tau \approx (k_3 + k_4)k_2/k_1 + k_4 = k_3(K_E + K_Q + K_E K_Q/[H^+]). \quad (9)$$

Relaxation times were measured from the slopes of the kinetic curves, i.e. of the logarithmic plots of excess optical densities (with respect to the equilibrium value) against time (Fig. 3). For pH 2.3 we obtained as mean values:

$$\begin{aligned} 1/\tau &= 5.12 (\pm 0.27) \times 10^{-3} \text{ sec}^{-1} \\ k_3 &= 4.05 (\pm 0.3) \times 10^6 \text{ M}^{-1} \text{ sec}^{-1} \\ k_4 &= k_3 K_Q = 7.9 (\pm 0.7) \times 10^{-4} \text{ sec}^{-1} \end{aligned}$$

DISCUSSION

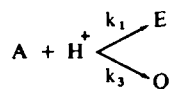
The tautomeric equilibrium of 3-TN in water favours the ketone form, Q: $K_T = 0.18$, whereas in ethanol the enol form, E, predominates: $K_T = 2.02$. The ratio of both constants is similar to many other keto-enol tautomeric systems.¹⁴ The enol form should predominate in non-polar solvents. Buu-Hoï *et al.*⁵ found that the NMR spectrum of 3-TN in CDCl_3 was that of the enol, E. We have found, however, that the NMR spectrum of 3-TN in CDCl_3 , the IR spectrum in CHCl_3 , and the UV spectrum in *n*-hexane are consistent with the pure ketone Q.

This apparent contradiction seems to arise from an extremely slow rate of tautomerization of crystalline 3-TN and of its solutions in non-polar solvents. The reaction lasts for months. Freshly sublimed 3-TN appears to be in pure ketone form. Our spectra were recorded just after dissolving freshly sublimed crystals. The sample of the cited authors⁵ most probably had been kept for sufficient time to approach the tautomeric equilibrium characterized by a large excess of enol.

The rate of tautomerization in aqueous solution, at 20°, is determined by a slow second step of reaction (1), according to Eq. (9). The rate constants found for proton transfer in this step are $k_3 = 4.05 \times 10^6 \text{ M}^{-1} \text{ sec}^{-1}$ and $k_4 = 7.9 \times 10^{-4} \text{ sec}^{-1}$. These values are consistent with the polarographic results of Zuman *et al.*⁹ who observed no kinetic character of the limiting currents of Q and A: the relaxation time in the anion-ketone system is much longer under their experimental conditions than the drop time in polarography.

The rate constants of proton transfer in the first step of reaction (1), i.e. in the anion-enol system, could not be determined by the rather slow method of measurements used by the present authors. However, both upper and lower limits are known

for these rate constants. On acidifying an alkaline solution of 3-TN both tautomers are formed in a competitive reaction



As E is formed in more than 95% , $k_1 \gg 20 k_3$. Another limit is given by the rate of diffusion:

$$10^8 \text{ M}^{-1} \text{ sec}^{-1} < k_1 < 10^{10} \text{ M}^{-1} \text{ sec}^{-1},$$

and, from $k_2 = k_1 K_E$:

$$10^{-1} \text{ sec}^{-1} < k_2 < 10 \text{ sec}^{-1}.$$

Although the protolytic equilibria for both systems are very close, a difference of 2 to 3 orders of magnitude is observed between the dissociation rate constants of ketone and of enol. This is a common difference between C-acids and O-acids.^{15, 16}

In numerous tautomeric systems, the absorption spectrum of a common anion resembles the spectrum of one of the tautomers.¹⁰ In the case of 3-TN, the UV spectrum of anion A (Fig. 1, curve c) is markedly different from both Q and E. As the tautomeric equilibrium constant K_T is of the order of 1, the electron density distribution in the anion appears to be intermediate between both "canonic formulas" of A, given in Eq. (1).

EXPERIMENTAL

3-Thianaphthenone was synthesized according to known procedure.¹⁷ Samples were sublimed under reduced press (about 10 Torr) at 75° before every series of measurements, and kept *in vacuo*.

All solns of 3-TN, especially alkaline solns, were thoroughly deaerated, as even traces of O₂ oxidize 3-TN to thioindigo which is insoluble in water. Oxygen was removed by bubbling a stream of N₂ purified on heated Cu through the solns. Aqueous solns of 3-TN, containing 5% EtOH, were prepared in the following way: a few crystals of 3-TN were placed in a dry and closed optical quartz cell, through which a stream of N₂ was passed. Then, 0.1 ml of deaerated EtOH was injected into the cell. After the crystals had dissolved, 1.9 ml of a suitable deaerated aqueous buffer soln was injected, and, after mixing and closing the stop-cocks, the spectra were measured.

For the kinetic measurements, 3-TN was dissolved in this manner in 2.2 ml of a borate buffer (0.007M Na₂B₄O₇ + 0.002M NaOH) of pH > 12, and—after the protolytic equilibrium has been established—0.5 ml of a 0.2M HCl was added to provide a final pH of 2.3. The soln was mixed by a stream of N₂ for 30 sec, and a first measurement of absorbance was made 60 sec after the injection of acid. After the reaction attained equilibrium, the concentration of 3-TN was determined from the optical density, using a formal absorption coefficient ϵ (370 nm) = 2500 M⁻¹cm⁻¹ (± 130), and a pH value was measured with a type PHM-4 pH-meter, Radiometer (Copenhagen).

Absorption spectra in UV were measured manually with a CF-4 spectrophotometer, Optica (Milan). Spectrometers type UR-10, C. Zeiss (Jena), and type MR-60, Varian (Palo Alto, Calif.) were used to record the IR and NMR spectra, respectively.

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