The acid dissociation constant of triphenylethenethiol, a simple thioenol, and that of its oxygen-enol analog

Yvonne Chiang, A. Jerry Kresge, Norman P. Schepp, Vladimir V. Popik, Zvi Rappoport, and Tzvia Selzer

Abstract: The acidity constant, $pQ_a^E = 8.49$, for the stable thioenol, triphenylethenethiol, was determined by spectrophotometric titration, and that, $pQ_a^E = 11.37$, for its unstable oxygen analog, triphenylethenol, was determined by analysis of its ketonization rate using enol generated flash photolytically from triphenylvinyl bromide. (These acidity constants are *concentration* quotients applicable at 0.10 M ionic strength in 50 vol.% aqueous methanol.) This appears to be the first determination of the acid strength of a simple thioenol. Triphenylethenethiol was unreactive in dilute acid or base, but in concentrated perchloric acid solutions it was slowly transformed into diphenylacetophenone and 2,3-diphenylbenzo[b]thiophene; the mechanisms of these reactions are believed to involve rapid equilibrium protonation of the enol on the β -carbon followed by rate-determining capture of the cation so formed, either externally by solvent or internally by one of the β -phenyl groups of the ion.

Key words: enol ketonization, thioenol, flash photolysis, vinyl halide.

Résumé : Utilisant un titrage spectrophotométrique, on a déterminé la constante d'acidité, pQ_a^E , du thioénol stable, triphényléthènethiol, alors que l'on a déterminé celle, pQ_a^E , de son analogue oxygéné instable, le triphényléthénol, par analyse de sa vitesse de cétonisation en utilisant l'énol généré par photolyse éclair à partir du bromure de triphénylvinyle. (Ces constantes d'acidité sont les quotients de *concentration* applicables à une force ionique de 0,10 M, dans une solution aqueuse de méthanol à 50% en volume.) Il semble que ce travail corresponde à la première détermination de la force acide d'un thioénol simple. Le triphényléthènethiol ne réagit pas en milieux acide ou base faibles; toutefois, en solutions concentrées d'acide perchlorique, il se transforme lentement en diphénylacétophénone et en 2,3-diphénylbenzo[*b*]thiophène. On croit que les mécanismes de ces réactions impliquent une protonation à équilibre rapide de l'énol sur le carbone- β , suivie d'une capture cinétiquement limitante du cation ainsi formé, d'une façon externe par le solvant ou externe par l'un des groupes β -phényles de l'ion.

Mots clés : cétonisation des énols, thioénol, photolyse éclair, halogénure de vinyle.

[Traduit par la rédaction]

Introduction

Simple enols are generally much less stable than their aldehyde or ketone tautomers (1). The same, however, is not true of simple thioenols: in the sulfur systems, enol and keto isomers usually exist in comparable amounts, and the enol is sometimes even the dominant tautomer (2). A recently examined example of the latter situation is provided by triphenylethenethiol, **1**, which could not be converted to its

Received December 16, 1997.

This paper is dedicated to Professor Erwin Buncel in recognition of his contributions to Canadian chemistry.

Y. Chiang, A.J. Kresge,¹ N.P. Schepp, and V.V. Popik. Department of Chemistry, University of Toronto, Toronto, ON M5S 3H6, Canada.

Z. Rappoport and T. Selzer. Department of Organic Chemistry, The Hebrew University, Jerusalem 91904, Israel.

Author to whom correspondence may be addressed. Telephone/Fax: (416) 978–7259. E-mail: akresge@alchemy.chem.utoronto.ca keto isomer, diphenylthioacetophenone, **2**, eq. [1], under conditions much more drastic than usually required to equilibrate oxygen keto–enol systems (3). This predominance of the enol isomer in the system of eq. [1] is consistent with the estimate $pK_E = 2.7$ for the oxygen system² and the suggestion that the



ratio of the keto–enol equilibrium constant of a sulfur system to that of its oxygen analog, $K_{\rm E}({\rm S})/K_{\rm E}({\rm O})$, will generally be equal to or greater than 10⁶ (3).

We have taken advantage of this stability of triphenylethenethiol to measure the equilibrium constant for its ionization as an acid, eq. [2], using a conventional spectroscopic method based upon the difference in uv absorbance of

² This estimate is based upon $pK_E = 0.98$ for diphenylacetaldehyde (4) and the effect of adding another phenyl group, obtained from $pK_E = 6.23$ for acetaldehyde (5) and $pK_E = 7.96$ for acetophenone (6).

658



the enol ($\lambda_{max} = 285$ nm) and its enolate ion, **3** ($\lambda_{max} = 315$ nm). This, to the best of our knowledge, is the first determination of the acidity constant of a simple thioenol.

To assess the effect on enol acidity of replacing oxygen with sulfur, we also determined the equilibrium constant for the ionization of triphenylethenol, **4**, eq. [3]. This enol, however, is not stable with respect to its keto isomer, and it reverts to the latter too rapidly to allow measurements by the method we employed for the thio analog. We therefore used a flash



photolytic technique instead: we generated the enol by photoionization of triphenylvinyl bromide, 5, and subsequent hydration of the ensuing vinyl cation, 6, eq. [4], and then determined the acidity constant of the enol, 4, from the dependence of its rate of ketonization to diphenylacetophenone, 7, upon the acidity of the medium.

Because some of the substrates used were insufficiently soluble in a purely aqueous solvent, this work was done in 50 vol.% aqueous methanol.



Experimental section

Materials

Triphenylethenethiol was a sample that had been prepared before (3), and 2,3-diphenylbenzo[*b*]thiophene was kindly provided by Dr. T. Kitamura (7). All other materials were best available commercial grades.

Spectrophotometric pK_a determination

The acid dissociation constant of triphenylethenethiol was determined by monitoring the absorbance of this substance in solutions of different acidity. The solvent was 50 vol.% aqueous methanol, and its acidity was controlled by using either perchloric acid or sodium hydroxide, or acetic acid, biphosphate ion, tris(hydroxymethyl)methylamine, or ammonia buffers; the ionic strength of these solutions was maintained at 0.10 M by adding sodium perchlorate as required.

Measurements were made at $\lambda = 340$ nm, the wavelength

which the difference in absorption between af triphenylethenethiol and its thiolate anion is greatest. A Cary 2200 spectrometer, with cell compartment thermostatted at 25.0 ± 0.05 °C, was used. The absorbance of the acid, base, or buffer solution without substrate was first recorded, a fixed amount of an acetonitrile stock solution of substrate (0.02 M) was then added (10.0 μ L to 3.00 mL or 5.00 μ L to 1.00 mL), and the absorbance was recorded again. The difference, ΔA , between such pairs of absorbance readings was then calculated, and the relationship between ΔA and the hydrogen ion concentration of the solutions was analyzed by least-squares fitting of eq. [5]; here $A_{\rm B}$ and $A_{\rm A}$ are the limiting absorbances of the basic and acidic forms of the substrate and Q is its acidity constant.3

[5]
$$\Delta A = \frac{A_{\rm B}Q_{\rm a}^{\rm E} + A_{\rm A}[{\rm H}^+]}{Q_{\rm a}^{\rm E} + [{\rm H}^+]}$$

Hydrogen ion concentrations of the solutions in which the absorbances were recorded were determined by measuring their pH using a Beckman Research pH meter, in conjunction with a ROSS combination electrode model no. 81-02 (Orion Research Inc.), and then converting that reading into [H⁺] with the aid of eq. [6] (8*a*).

[6]
$$-\log[H^+] = pH_{meas} + \log \gamma_{H^+} - \delta = pH_{meas} - 0.205$$

In this equation, γ_{H^+} is the activity coefficient of H⁺, which was calculated using the Debye–Hückel equation (8*b*) with parameters appropriate to the medium employed, and δ is an empirical parameter (8*a*). The validity of this method was checked by measuring the pH of five perchloric acid solutions of known concentration in the range 0.02–0.10 M. The results obtained gave an average value of $(-\log[\text{H}^+] - \text{pH}_{\text{meas}}) = -0.2049 \pm 0.0061$, in excellent agreement with eq. [6].

Flash photolysis

Flash photolytic rate determinations were made using a conventional flash-lamp system with cell compartment thermostatted at 25.0 ± 0.05 °C that has already been described (5). Measurements were carried out in perchloric acid and sodium hydroxide solutions whose ionic strength was maintained at 0.10 M; substrate (triphenylvinyl bromide) concentrations were $(2-3) \times 10^{-5}$ M. The lifetimes of the transient species observed in basic solution were short enough to be analyzed by the flash photolysis monitoring system, but those observed in acidic solution were too long-lived (t = 15-40 min). For rate measurements in acidic solutions, therefore, transients were generated by a single flash in the flash system, and the reacting solutions were then transferred to a Cary 118 spectrometer, with cell compartment thermostatted at 25.0 ± 0.05 °C, for rate measurement there. Decay of the transients observed in both kinds of experiment obeyed the first-order rate law accurately, and observed first-order rate constants were calculated by least-squares fitting of an exponential function.

Results

Spectrophotometric pK_a determination

The equilibrium constant for ionization of triphenylethenethiol

³ This is a *concentration* dissociation constant, applicable at ionic strength = 0.10 M.

Fig. 1. Spectrophotometric titration curve for the acid ionization of triphenylethenethiol in 50 vol.% aqueous methanol at 25°C.



as an acid, eq. [2], was determined by measuring the absorbance of this substance dissolved in 50 vol.% aqueous methanol solutions of different acidity. Two series of experiments were carried out, one at a fixed total concentration of substrate of 7×10^{-5} M and another at 1×10^{-4} M. The data so obtained are summarized in Table S1.⁴

As Fig. 1 illustrates, these data describe the expected titration curve, eq. [5], well. Least-squares analysis gave the results $Q_a^E = (3.48 \pm 0.22) \times 10^{-9}$ M, $pQ_a^E = 8.46 \pm 0.03$, and $Q_a^E = (3.07 \pm 0.21) \times 10^{-9}$ M, $pQ_a^E = 8.51 \pm 0.03$. The weighted average of these gives $Q_a^E = (3.27 \pm 0.15) \times 10^{-9}$ M, $pQ_a^E = 8.49 \pm 0.02$, as the best value of this acidity constant.

Reaction of triphenylethenethiol in strong acid

Except for the acid ionization described above, triphenylethenethiol proved to be unreactive in dilute 50 vol.% aqueous methanol solutions of acid or base. In concentrated perchloric acid, however, its uv absorbance did decrease slowly in an acid-catalyzed process, and some rate measurements were made by monitoring this decay at $\lambda = 300$ nm. The rate constants so obtained are summarized in Table S2.⁴ These data were extrapolated down to dilute solution by the Cox–Yates method (9) using the X_0 acidity function (10) on the assumption that values of X_0 in 50 vol.% aqueous methanol are the same as those in wholly aqueous solutions of the same acid concentration. This gave $k_{\rm H^+} = (2.72 \pm 0.41) \times 10^{-8} \, {\rm M^{-1} \, s^{-1}}$ as the dilute solution rate constant.

A product study was performed by allowing 0.0025 M triphenylethenethiol dissolved in 4 M perchloric acid in 2:1 methanol:water (vol/vol) to react for 36 h, and then subjecting the resulting solution to HPLC analysis. This showed that two products were formed in comparable amounts, and spiking with authentic samples showed them to be diphenylacetophenone, **7**, and 2,3-diphenylbenzo[*b*]thiophene, **8**.

Fig. 2. Relationship between sodium hydroxide concentration and rates of ketonization of triphenylethenol in 50 vol.% aqueous methanol solution at 25°C.



Flash photolysis

Flash photolysis of vinyl halides is known to produce vinyl cations (11), which, in the presence of water, are hydrated to enols; the enols then ketonize to their carbonyl isomers (11b,c). Evidence that this is the sequence of reactions that occurred in the present study upon flash photolysis of triphenylvinyl bromide, as depicted in eq. [4], is provided by the uv spectral changes that took place. The uv spectrum of triphenylvinyl bromide in 50 vol.% aqueous methanol consists of a strong absorption band with $\lambda_{max} = 290$ nm. Upon flashing, this band immediately increased in intensity in a process that was complete within the duration of the flash (50 ms). Such a change in intensity without the appearance of any new absorption bands is to be expected for the conversion of a vinyl halide to a vinyl alcohol with no change in the other doublebond substituents; the lifetime of vinyl cations in aqueous media, moreover, is known to be much shorter than 50 ms (11c), and formation of the enol would be expected to be immediate on the time scale of the present experiment. This increase in absorbance at $\lambda = 290$ nm was followed by a decrease whose velocity was catalyzed by both hydronium and hydroxide ions, with saturation of the hydroxide ion catalysis at high basicities. Such behavior is characteristic of the ketonization of simple enols, where saturation of hydroxide ion catalysis represents conversion of enol to enolate ion (12). The absorbance decrease at $\lambda = 290$ was also accompanied by an increase occurring at the same rate at $\lambda = 250$ nm, where diphenylacetophenone, the expected product of the present ketonization reaction, has an absorption maximum. It appears safe to conclude, therefore, that the process monitored here was ketonization of triphenylethenol to diphenylacetophenone.

Rates of this ketonization reaction were measured in

Tables S1–S4 may be purchased from: The Depository of Unpublished Data, Document Delivery, CISTI, National Research Council Canada, Ottawa, Canada K1A OS2.

0.02–0.10 M perchloric acid solutions. The data are summarized in Table S3.⁴ Observed rate constants proved to be accurately proportional to acid concentration, and the data were therefore analyzed by least-squares fitting of eq. [7]. This produced the result $k_{\rm H^+} = (1.05 \pm 0.04) \times 10^{-2} \, {\rm M^{-1} \ s^{-1}}$ for the acid-catalyzed reaction and $k_{\rm uc} = (2.11 \pm 0.30) \times 10^{-4} \, {\rm s^{-1}}$ for an "uncatalyzed" process.

[7]
$$k_{\rm obs} = k_{\rm uc} + k_{\rm H^+} [\rm H^+]$$

Rates of ketonization of triphenylethenol were also measured in sodium hydroxide solutions over the concentration range 0.002–0.10 M. These data are summarized in Table S4.⁴ As Fig. 2 shows, observed first-order rate constants increased with increasing basicity at low hydroxide ion concentrations, but at higher basicity the effect diminished, and rate constants finally reached a constant value independent of hydroxide ion concentration. Such behavior is commonly found in enol ketonization. It is known to be caused by a shift in the initial state of the reaction from enol to enolate ion, plus the facts that ketonization occurs by rate-determining protonation of the substrate on the β -carbon and that enolate ions are much more susceptible to such electrophilic attack than are enols (12). In weakly basic solutions, therefore, where the substrate exists mainly as un-ionized enol, ketonization will occur as shown in eq. [8]: rapid equilibrium ionization of the enol will be followed by rate-determining reaction of the more labile enolate ion by proton transfer from the only available proton donor, the hydroxylic solvent. Since hydrogen ions are produced in the pre-equilibrium step but are not used up in the rate-determining step, the overall reaction rate will show an inverse dependence on hydrogen ion concentration or a direct dependence on hydroxide ion concentration. This apparent hydroxide ion catalysis, however, will diminish as the hydroxide ion concentration increases and the position of the pre-equilibrium step



shifts toward enolate ion, and that will produce the saturation behavior shown in Fig. 2.

The rate law that applies to the reaction scheme of eq. [8] is shown in eq. [9]. Least-squares fitting of the data using this equation gave $k'_{o} = 63.1 \pm 0.07 \text{ s}^{-1}$ and $Q_a^{E} = (4.28 \pm 0.22) \times 10^{-12} \text{ M}$, $pQ_a^{E} = 11.37 \pm 0.02.^3$ Hydrogen ion concentrations needed for this purpose were obtained from pH meter readings with the aid of eq. [6].

[9]
$$k_{\rm obs} = \frac{k'_{\rm o} Q_{\rm a}^{\rm E}}{[{\rm H}^+] + Q_{\rm a}^{\rm E}}$$

Discussion



Comparison of the rate constants obtained here for the acidcatalyzed ketonization of triphenylethenol and the disappearance of triphenylethenethiol in concentrated perchloric acid solutions shows that the oxygen enol is 3.8×10^5 times as reactive as the thioenol. This is an unusually large oxygen-tosulfur rate ratio for rate-determining protonation of a carbon–carbon double bond activated by these atoms; for example, the acid-catalyzed hydrolysis of ethyl vinyl ether, **9**, and ethyl vinyl sulfide, **10**, both of which occur by rate-determining protonation of the substrate on its β -carbon atom, give an oxygento-sulfur rate ratio of only 37 (13), some 10 000 times smaller than the present value.

This striking difference suggests that the disappearance of triphenylethenethiol in concentrated perchloric acid does not represent rate-determining protonation of this substance on its β -carbon atom, but rather that such protonation occurs rapidly and reversibly and is then followed by some other, much slower, reaction step. This hypothesis is consistent with the conclusion reached before (3) that the enol in this system is more stable than the thicketone, for the ketone, once formed, would then revert to the enol. It is also consistent with our observation that the oxygen ketone and a thiophene derivative, rather than the thicketone, are the products of the concentrated acid reaction. It is likely that the oxygen ketone is formed by occasional capture, by water or methanol, of the cation produced by reversible carbon protonation of the thioenol, followed by decomposition of that adduct, as shown in eq. [10], and that the thiophene product is formed by intramolecular capture of this cation by one of its β -phenyl groups, in an electrophilic aromatic substitution reaction, followed by aromatization of the thiophene ring, as shown in eq. [11]; a similar cyclization to a benzofuran has been observed in an analogous stable oxygen-enol system (14).

These observations reinforce the conclusion reached before (3) that the enol is the more stable tautomer in the diphenylthioacetophenone keto–enol system. Additional evidence that this is so is provided by the fact that we were able to determine the acidity constant of this thioenol by a conventional slow method. The limiting rate constant for ketonization of the corresponding oxygen enol, triphenylethenol, determined by our flash photolysis experiments, is 63 s^{-1} . Applying a sulfur-to-oxygen rate ratio of ca. 10^{-2} to this, as expected, for example, on the basis of the rates of reaction of ethyl vinyl





ether and ethyl vinyl sulfide discussed above, puts the lifetime of triphenylethenethiol at about 1 s; this short lifetime would have made our conventional measurements impossible if the enol were not the stable isomer.

We are aware of no previous determination of the acidity constant of a simple thioenol, and we consequently cannot compare the value we have obtained with other results. It is interesting. however, that our value makes triphenylethenethiol, with $pQ_a^E = 8.49$, a considerably stronger acid than its oxygen analog, triphenylethenol, with $pQ_a^{\bar{E}} =$ 11.37, for this is consistent with the fact that simple thiols are considerably stronger acids than simple oxygen alcohols (15). A quantitative comparison might be more appropriately based upon phenols rather than alcohols, for phenols, like enols, have their acidic groups attached to carbon-carbon double bonds, and enol pK_a 's are in fact quite similar to phenol pK_a 's (12). It is significant, therefore, that the difference in acid strength between the oxygen and sulfur enois examined here, $\Delta p Q_a^E$ = 2.88, is not unlike the difference in acidity between phenol $(pK_a = 10.00 (16))$ and thiophenol $(pK_a = 6.49 (17)), \Delta pK_a =$ 3.51.

Acknowledgements

We are grateful to Professor T. Kitamura for supplying 2,3diphenylbenzo[*b*]thiophene and to the Natural Sciences and Engineering Research Council of Canada for financial support of this research.

References

- 1. J. Toullec. *In* The chemistry of enols. *Edited by* Z. Rappoport. Wiley–Interscience, New York. 1990. pp. 323–398.
- F. Duus. *In* Comprehensive organic chemistry. Vol. 3. *Edited by* D. Barton and W.D. Ollis. Pergamon Press, Exeter, U K. 1979. pp. 385–388.

- 3. T. Selzer and Z. Rappoport. J. Org. Chem. 61, 5462 (1996).
- Y. Chiang, A.J. Kresge, and E.T. Krogh. J. Am. Chem. Soc. 110, 2600 (1988).
- Y. Chiang, M. Hojatti, J.R. Keeffe, A.J. Kresge, N.P. Schepp, and J. Wirz. J. Am. Chem. Soc. 109, 4000 (1987).
- J.R. Keeffe, A.J. Kresge, and J. Toullec. Can. J. Chem. 64, 1224 (1986).
- 7. T. Kitamura, S. Soda, K. Morizane, Y. Fujiwara, and H. Taniguchi. J. Chem. Soc. Perkin Trans. 2, 473 (1996).
- 8. R.G. Bates. Determination of pH theory and practice. Wiley–Interscience, New York. 1973. (*a*) pp. 243–245; (*b*) p. 216.
- 9. R.A. Cox. Acc. Chem. Res. 20, 27 (1987).
- R.A.Cox and K. Yates. Can. J. Chem. **59**, 2116 (1981);
 A.J. Kresge, H.J. Chen, G.L. Capen, and M.F. Powell. Can. J. Chem. **61**, 249 (1983).
- (a) W. Schnabel, I. Naito, T. Kitamura, S. Kobayashi, and H. Taniguchi. Tetrahedron, **36**, 3229 (1980); S. Kobayashi, T. Kitamura, H. Taniguchi, and W. Schnabel. Chem. Lett. 2101 (1983); S. Kobayashi, Q.Q. Zhu, and W. Schnabel. Z. Naturforsch. B: Chem. Sci. **43**, 825 (1988); F.I.M. van Ginkel, R.J. Wisser, C.A.G.O. Varma, and G. Lodder. J. Photochem. **30**, 453 (1985); R.A. McClelland, F. Cozens, and S. Steenken. Tetrahedron Lett. **31**, 2821 (1990); (b) A.J. Kresge and N.P. Schepp. J. Chem. Soc. Chem. Commun. 1548 (1989); (c) Y. Chiang, R. Eliason, J. Jones, A.J. Kresge, K.L. Evans, and R.D. Gandour. Can. J. Chem. **71**, 1964 (1993).
- J.R. Keeffe and A.J. Kresge. *In* The chemistry of enols. *Edited* by Z. Rappoport. Wiley–Interscience, New York. 1990. pp. 399–480.
- 13. R.A.McClelland. Can. J. Chem. 55, 548 (1977).
- M. Schmittel and U Baumann. Angew. Chem. Int. Ed. Engl. 29, 541 (1990).
- 15. R. Stewart. The proton: applications to organic chemistry. Academic Press, New York. 1985. pp. 45–46.
- 16. A.J. Biggs. Trans. Faraday Soc. 52, 35 (1956).
- C.L. Liotta, E.M. Perdue, and H.P. Hopkins, Jr. J. Am. Chem. Soc. 96, 7981 (1974).