

# Radical Cations in the OH-Radical-Induced Oxidation of Thiourea and Tetramethylthiourea in Aqueous Solution

Wen-feng Wang,<sup>†,§</sup> Man Nien Schuchmann,<sup>†</sup> Heinz-Peter Schuchmann,<sup>†</sup> Wolfgang Knolle,<sup>‡</sup> Justus von Sonntag,<sup>‡</sup> and Clemens von Sonntag<sup>\*,†</sup>

Contribution from the Max-Planck-Institut für Strahlenchemie, Stiftstrasse 34-36, P.O. Box 101365, D-45413 Mülheim an der Ruhr, Germany, Institut für Oberflächenmodifizierung, Permoserstrasse 15, D-04303 Leipzig, Germany, and Shanghai Institute for Nuclear Research, P.O. Box 800-204, Shanghai 201800, China

Received September 14, 1998

**Abstract:** Hydroxyl radicals were generated radiolytically in N<sub>2</sub>O-saturated aqueous solutions of thiourea and tetramethylthiourea. The rate constant of the reaction of OH radicals with thiourea (tetramethylthiourea) has been determined using 2-propanol as well as NaN<sub>3</sub> as competitors to be  $1.2 \times 10^{10} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  ( $8.0 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ). A transient appears after a short induction period and shows a well-defined absorption spectrum with  $\lambda_{\text{max}} = 400 \text{ nm}$  ( $\epsilon = 7400 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}$ ); that of tetramethylthiourea has  $\lambda_{\text{max}} = 450 \text{ nm}$  ( $\epsilon = 6560 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}$ ). Using conductometric detection, it has been shown that, in both cases, OH<sup>-</sup> and a positively charged species are produced. These results indicate that a radical cation is formed. These intermediates with  $\lambda_{\text{max}} = 400 \text{ nm}$  (450 nm) are not the primary radical cations, since the intensity of the absorbance depends on the substrate concentration. The absorbance build-up follows a complex kinetics best described by the reversible formation of a dimeric radical cation by addition of a primary radical cation to a molecule of thiourea. The equilibrium constant for this addition has been determined by competition kinetics to be  $5.5 \times 10^5 \text{ dm}^3 \text{ mol}^{-1}$  for thiourea ( $7.6 \times 10^4 \text{ dm}^3 \text{ mol}^{-1}$  for tetramethylthiourea). In the bimolecular decay of the dimeric radical cation (thiourea,  $2k = 9.0 \times 10^8 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ; tetramethylthiourea,  $1.3 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ), formamidine (tetramethylformamidine) disulfide is formed. In basic solutions of thiourea, the absorbance at 400 nm of the dimeric radical cation decays rapidly, giving rise ( $5.9 \times 10^7 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ) to a new intermediate with a broad maximum at 510 nm ( $\epsilon = 750 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}$ ). This reaction is not observed in tetramethylthiourea. The absorption at 510 nm is attributed to the formation of a dimeric radical anion, via neutralization of the dimeric radical cation and subsequent deprotonation of the neutral dimeric radical. The primary radical cation of thiourea is deprotonated by OH<sup>-</sup> ( $2.8 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ) to give a neutral thiyl radical. The latter reacts rapidly with thiourea, yielding a dimeric radical, which is identical to the species from the reaction of OH<sup>-</sup> with the dimeric radical cation. The dimeric radical cations of thiourea and tetramethylthiourea are strong oxidants and readily oxidize the superoxide radical ( $4.5 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  for thiourea and  $3.8 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  for tetramethylthiourea), phenolate ion ( $3 \times 10^8 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  for tetramethylthiourea), and even azide ion ( $4 \times 10^6 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  for thiourea and  $\sim 10^6 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  for tetramethylthiourea). With O<sub>2</sub>, the dimeric radical cation of thiourea reacts relatively slowly ( $1.2 \times 10^7 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ) and reversibly ( $2 \times 10^3 \text{ s}^{-1}$ ).

## Introduction

Sulfur-containing compounds show a complex free-radical chemistry which is very different from that exhibited by carbon-centered free radicals (for reviews see refs 1–4). For example, sulfur-centered radicals are able to form three-electron-bonded intermediates such as RSSR<sup>-</sup>, R<sub>2</sub>SSR<sub>2</sub><sup>•+</sup>, or R<sub>2</sub>SOH<sup>•</sup>.<sup>5</sup> The radiation-induced chemistry of thiourea in aqueous solution has been studied<sup>6–13</sup> with a view toward its potential application as

a radioprotectant. Formamidine disulfide appears to be a primary product; several sulfur-free nitrogen-containing compounds, sulfate ion, and elemental sulfur have also been observed. These

\* To whom correspondence should be addressed. Fax: (Germany) 208 306 3951. Phone: (Germany) 208 306 3529. E-mail: vonsonntag@mpi-muelheim.mpg.de.

<sup>†</sup> Max-Planck-Institut für Strahlenchemie.

<sup>‡</sup> Institut für Oberflächenmodifizierung.

<sup>§</sup> Shanghai Institute for Nuclear Research.

(1) Asmus, K.-D. *Acc. Chem. Res.* **1979**, *12*, 436–442.

(2) von Sonntag, C.; Schuchmann, H.-P. In *The Chemistry of Functional Groups. Supplement E: The Chemistry of Ethers, Crown Ethers, Hydroxyl Groups and Their Sulphur Analogues. Part 2*; Patai, S., Ed.; Wiley: New York, 1980; pp 971–993.

(3) Asmus, K.-D. *Methods Enzymol.* **1990**, *186*, 168–180.

(4) Wardman, P.; von Sonntag, C. *Methods Enzymol.* **1995**, *251*, 31–45.

(5) Asmus, K.-D. In *Sulfur-Centered Reactive Intermediates in Chemistry and Biology*; Chatgililoglu, C., Asmus, K.-D., Eds.; Plenum: New York, 1990; pp 155–172.

(6) Dale, W. M. *Nature* **1956**, *177*, 531.

(7) Dale, W. M.; Davies, J. V. *Progress in Radiobiology*, Proceedings, 4th Int. Conf., Cambridge, 1955; pp 119–127.

(8) Dale, W. M.; Davies, J. V. *Radiat. Res.* **1957**, *7*, 35–46.

(9) Dale, W. M.; Davies, J. V. *Int. J. Radiat. Biol.* **1959**, *1*, 189–195.

(10) Charlesby, A.; Kopp, P. M.; Read, J. F. *Proc. R. Soc. A* **1966**, *292*, 122–133.

(11) Nanobashvili, E. M.; Chirakadse, G. G.; Gvilava, S. E.; Mosashvili, G. A.; Katsadse, D. V. VI. *Thiourea, urea and their derivatives*; Izdatel'stvo "Metsniereba": Tbilisi, 1980; pp 113–136.

(12) Natsvlshvili, S. E.; Nanobashvili, H. M.; Ignatashvili, S. H.; Panchvidze, M. V. In *Proceedings of the 5th Tihany Symp. Radiation Chemistry*; Dobo, J., Hedvig, P., Schiller, R., Eds.; Akademiai Kiado: Budapest, 1983; pp 589–593.

(13) Charlesby, A.; Fydeler, P. J.; Kopp, P. M.; Keene, J. P.; Swallow, A. J. In *Pulse Radiolysis*; Ebert, M., Keene, J. P., Swallow, A. J., Baxendale, J. H. Eds.; Academic Press: London, 1965; pp 193–201.

latter products are probably, in most cases, largely hydrolysis products of the disulfide (cf. ref 14), but there are also indications that, in the presence of oxygen, a chain process may give rise to the production of elemental sulfur under certain conditions.<sup>8</sup> Recently, the radiolysis of thiourea derivatives has been studied in connection with their use as corrosion inhibitors in nuclear reactors.<sup>15,16</sup>

The reactions of OH radicals with thiourea derivatives give rise to an intermediate which exhibits a characteristic absorption near 400 nm.<sup>13,15,16</sup> While this intermediate has, in the past, been thought to be a neutral radical, the present work will show that, with thiourea and *N,N,N',N'*-tetramethylthiourea, the intermediate in question is, in fact, a dimeric radical cation.

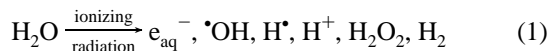
## Experimental Section

Pulse radiolysis of millimolar solutions of thiourea (Merck) and *N,N,N',N'*-tetramethyl thiourea (Fluka) in Milli-Q-filtered (Millipore) water saturated with either N<sub>2</sub>O or with N<sub>2</sub>O/O<sub>2</sub> was carried out (in Mülheim) with a 2.8-MeV Van-de-Graaff generator delivering electron pulses of 0.4–2 μs. Intermediates were monitored by optical and conductometric detection. The pulse radiolysis setup has been described previously.<sup>17</sup> Nanosecond-pulse experiments were carried out (in Leipzig) with an 11 MeV linear accelerator, with 17-Gy pulses of 5-ns duration. For dosimetry, N<sub>2</sub>O-saturated thiocyanate solution was used for optical detection,<sup>18</sup> and dimethyl sulfoxide solution was used for conductometric detection.<sup>19</sup> A series of pulse radiolysis experiments was also carried out in the presence of varying concentrations of the competitors 2-propanol and sodium azide.

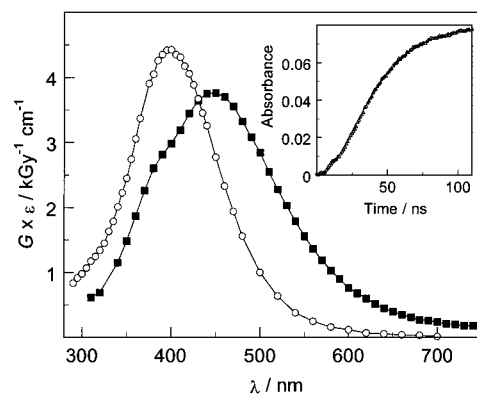
As the expected<sup>10,20</sup> primary oxidation products of the radiolysis of the thioureas, the formamidine and tetramethylformamidine disulfides were synthesized<sup>21,22</sup> from the corresponding thioureas through oxidation with H<sub>2</sub>O<sub>2</sub>, to be used as reference material. This procedure yields the formamidine disulfides, which are stable under acidic conditions. Their spectra turn out to be quite similar to those of their parent thioureas. The fact that we have not succeeded in separating these disulfides from their parent thioureas by HPLC has, however, precluded their direct identification and quantitative determination as products in the present case.

## Results and Discussion

**Radical-Generating System.** Hydroxyl radicals are generated in the radiolysis of water (reaction 1). The radiation chemical yields (*G* values) of the primary radicals are  $G(\bullet\text{OH}) \approx G(e_{\text{aq}}^-) = 2.9 \times 10^{-7} \text{ mol J}^{-1}$  and  $G(\text{H}\bullet) = 0.6 \times 10^{-7} \text{ mol J}^{-1}$ . N<sub>2</sub>O is used to convert the solvated electron into  $\bullet\text{OH}$  (reaction 2) so that a total hydroxyl radical yield of  $5.8 \times 10^{-7} \text{ mol J}^{-1}$  is achieved in a  $10^{-3} \text{ mol dm}^{-3}$  thiourea solution.  $G(\bullet\text{OH})$  increases with increasing substrate concentration.<sup>23</sup>



The reaction rates of the OH radical, the solvated electron, and the H atom with thiourea are close to the diffusion-controlled limit<sup>24</sup> (the value reported in ref 24 for  $k(\bullet\text{OH} +$



**Figure 1.** Spectrum (2 μs after the pulse) obtained in the pulse radiolysis of N<sub>2</sub>O-saturated solutions of thiourea (O,  $2 \times 10^{-3} \text{ mol dm}^{-3}$ ) and tetramethylthiourea (■,  $10^{-3} \text{ mol dm}^{-3}$ ) at natural pH (6.8); 3 Gy/pulse. Inset: Pulse radiolysis (5 ns, 17 Gy/pulse) of N<sub>2</sub>O-saturated solution of thiourea ( $10^{-2} \text{ mol dm}^{-3}$ ) at pH 6.5. Absorption build-up monitored at 400 nm (Δ). The solid line is the result of computational modeling of the reactions 3–5 (see text).

thiourea),  $3.9 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ , obtained by monitoring the build-up kinetics of the intermediate at 400 nm, is too low, as will be shown below).

**Reaction of the OH Radical with Thiourea and Tetramethylthiourea.** Pulse radiolysis of N<sub>2</sub>O-saturated solution of thiourea yields an intermediate which exhibits a strong absorption with  $\lambda_{\text{max}} = 400 \text{ nm}$  (Figure 1). Taking  $G(\bullet\text{OH}) = 5.96 \times 10^{-7} \text{ mol J}^{-1}$ ,  $\epsilon(400 \text{ nm}) = 7400 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}$  is calculated. The spectrum of the intermediate derived from tetramethylthiourea has its maximum at a longer wavelength [ $\epsilon(450 \text{ nm}) = 6560 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}$ , Figure 1]. This is comparable to the case of *N,N'*-diethylthiourea [ $\epsilon(425 \text{ nm}) = 5250 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}$ ].<sup>16</sup> The similarity that is observed in the behavior of thiourea and tetramethylthiourea toward the OH radical indicates that this reaction is not an H atom abstraction. (Thiourea had formerly often been thought to exist in the iso, i.e., iminothiol, form, cf. ref 25, but this view has since been refuted.<sup>26</sup>)

A complexity is apparent in the way the absorption builds up at 400 nm. While its rate increases at higher thiourea concentrations, the build-up shows a lag period of several nanoseconds before it takes on first-order characteristics (Figure 1, inset).

The build-up kinetics as shown in Figure 1 (inset) and the other findings described below are in accordance with the sequence of reactions 3–5 (Scheme 1; R = H, thiourea; R = CH<sub>3</sub>, tetramethylthiourea).

OH radical addition could occur initially at the carbon–sulfur double bond (reaction 3). The short-lived adduct could then decay into a radical cation and a hydroxide ion (reaction 4). The existence of the lag phase in the absorbance build-up (Figure 1, inset) indicates an initial reaction which gives rise to an intermediate which does not absorb at 400 nm.

At the same time, conductance measurements show that, with both thiourea and tetramethylthiourea, a positively charged

(14) Rio, L. G.; Munkley, C. G.; Stedman, G. *J. Chem. Soc., Perkin Trans. 2*, **1996**, 159–162.

(15) Dey, G. R.; Naik, D. B.; Kishore, K.; Moorthy, P. N. *Radiat. Phys. Chem.* **1994**, *43*, 365–369.

(16) Dey, G. R.; Naik, D. B.; Kishore, K.; Moorthy, P. N. *J. Chem. Soc., Perkin Trans. 2* **1994**, 1625–1629.

(17) von Sonntag, C.; Schuchmann, H.-P. *Methods Enzymol.* **1994**, *233*, 3–20.

(18) Schuler, R. H.; Patterson, L. K.; Janata, E. *J. Phys. Chem.* **1980**, *84*, 2088–2089.

(19) Schuchmann, H.-P.; Deeble, D. J.; Phillips, G. O.; von Sonntag, C. *Radiat. Phys. Chem.* **1991**, *37*, 157–160.

(20) Simonet, J. In *The Chemistry of Sulfur-Containing Functional Groups*; Patai, S., Rappoport, Z., Eds.; Wiley: New York, 1993; Suppl. S., pp 440–492.

(21) Preisler, P. W.; Berger, L. *J. Am. Chem. Soc.* **1947**, *69*, 322–325.

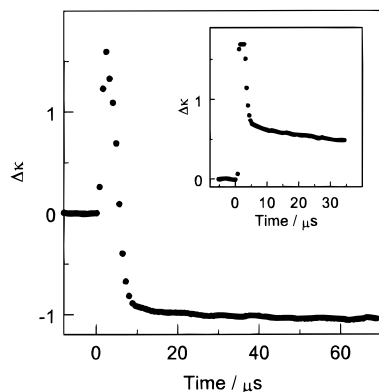
(22) Foss, O.; Johnsen, J.; Tvedten, O. *Acta Chem. Scand.* **1958**, *12*, 1782–1798.

(23) Schuler, R. H.; Hartzell, A. L.; Behar, B. *J. Phys. Chem.* **1981**, *85*, 192–199.

(24) Buxton, G. V.; Greenstock, C. L.; Helman, W. P.; Ross, A. B. *J. Phys. Chem. Ref. Data* **1988**, *17*, 513–886.

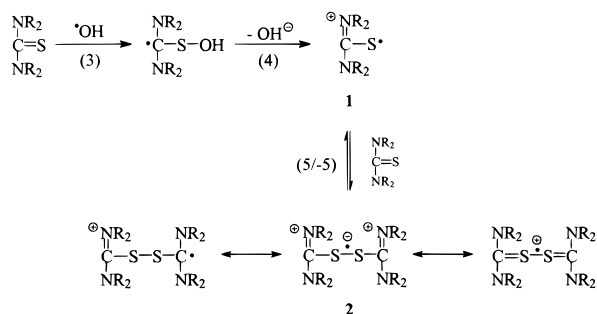
(25) Werner, E. A. *J. Chem. Soc.* **1912**, 2180–2191.

(26) Duss, F. In *Comprehensive Organic Chemistry*; Neville Jones, D., Ed.; Pergamon Press: Oxford, 1979; pp 373–487.



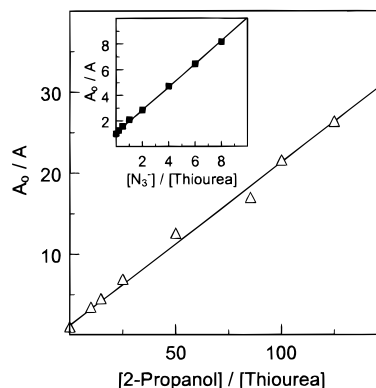
**Figure 2.** Conductance change observed in  $N_2O$ -saturated solution of thiourea ( $10^{-3} \text{ mol dm}^{-3}$ ) at pH 3.5 and at pH 10 (inset);  $0.4\text{-}\mu\text{s}$  pulses of  $\sim 4 \text{ Gy/pulse}$ . The initial jump at  $\leq 10 \mu\text{s}$  is due to the formation and subsequent neutralization of  $H^+$  within the pulse (reaction 1).

### Scheme 1



species is generated. In acidic solutions of thiourea, a conductance decrease is observed immediately ( $< 5 \mu\text{s}$ ) after the pulse (Figure 2), whereas in basic solution, a conductance build-up is observed (Figure 2, inset). This indicates that the formation of an ion pair (reaction 4) is the quasi-immediate consequence of the reaction of the hydroxyl radical. In acidic solution, the hydroxide anion formed in reaction 4 removes  $H^+$ , which leads to a decrease of the conductance. In basic solution, the ion pair from reaction 4 remains unaffected, which results in a conductance increase.

As the bimolecular rate constant of reaction 3 cannot be determined directly from the build-up kinetics at 400 nm (as has been attempted elsewhere),<sup>15,16,24</sup> we have determined it by competition with 2-propanol. In  $N_2O$ -saturated solutions of 2-propanol containing various concentrations of thiourea ( $4 \times 10^{-4}$ – $5 \times 10^{-3} \text{ mol dm}^{-3}$ ), the absorbance ( $A$ ) at 400 nm, 10  $\mu\text{s}$  after the pulse, was determined at each thiourea concentration ( $A_0$  is the absorbance of the thiourea solution at the same wavelength in the absence of 2-propanol). Taking  $k(\cdot\text{OH} + 2\text{-propanol}) = 1.9 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ,<sup>24</sup>  $k_3(\cdot\text{OH} + \text{thiourea}) = 1.0 \times 10^{10} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  is obtained from the slope of the plot  $A_0/A$  shown in Figure 3. Similarly,  $k_3(\cdot\text{OH} + \text{tetramethylthiourea}) = 8.0 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  has been determined. We have also determined  $k_3(\cdot\text{OH} + \text{thiourea})$  by competition with sodium azide (Figure 3, inset). This is possible since the reverse reaction between  $N_3^\ominus$  and thiourea does not materialize (see below). Taking  $k(\cdot\text{OH} + N_3^\ominus) = 1.2 \times 10^{10} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ,<sup>24</sup>  $k_3(\cdot\text{OH} + \text{thiourea}) = 1.3 \times 10^{10} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  is obtained from the slope of the plot shown in the Figure 3 inset. From these two independent measurements, the average value for  $k_3(\cdot\text{OH} + \text{thiourea})$  is  $1.2 \times 10^{10} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  (for a compilation of the rate constants determined in this study, see Table 1). The thiourea concentrations used in these competition



**Figure 3.** Competition plot of hydroxyl radical reactions with thiourea ( $4 \times 10^{-4}$ – $5 \times 10^{-3} \text{ mol dm}^{-3}$ ) and with 2-propanol ( $0.05 \text{ mol dm}^{-3}$ ). Inset: Competition plot of hydroxyl radical reactions with thiourea ( $10^{-3} \text{ mol dm}^{-3}$ ) and with  $NaN_3$  ( $2 \times 10^{-4}$ – $8 \times 10^{-3} \text{ mol dm}^{-3}$ ).

experiments ensured (see below) that equilibrium 5/–5 lay practically completely on the side of **2**.

Taking the equivalent conductance at 20 °C of  $H^+$  ( $325 \Omega^{-1} \text{ cm}^3 \text{ mol}^{-1}$ ),  $OH^-$  ( $180 \Omega^{-1} \text{ cm}^3 \text{ mol}^{-1}$ )<sup>27</sup> into account, and assuming a value of  $50 \Omega^{-1} \text{ cm}^3 \text{ mol}^{-1}$  for the singly charged radical cation, one obtains  $G(\text{radical cation}) = (5.8 \pm 0.5) \times 10^{-7} \text{ mol J}^{-1}$  (limit at relatively high thiourea concentrations, e.g.,  $10^{-3} \text{ M}$ ). This confirms that the  $OH^\cdot$  radical reaction leads essentially quantitatively to radical cation formation. Additional support for this has been obtained as follows. Oxidation of thiourea with the strong oxidant  $Cl_2^{\cdot\ominus}$  in  $0.1 \text{ mol dm}^{-3} NaCl$  and  $10^{-3} \text{ mol dm}^{-3}$  thiourea solution produces the radical cation of thiourea quantitatively since  $Cl_2^{\cdot\ominus}$  is a poor hydrogen atom abstractor. By comparing the yield of the radical cation obtained in this way with that obtained by the oxidation with  $\cdot\text{OH}$ , we found identical values. Thus, the oxidation of thiourea by  $\cdot\text{OH}$  produces the radical cation exclusively.

With tetramethylthiourea, oxidation by  $\cdot\text{OH}$  produces only 85% of radical cation as compared with the oxidation by  $Cl_2^{\cdot\ominus}$  (measured by optical detection). This implies that, in this case, there may be some 15% H atom abstraction by  $\cdot\text{OH}$  at the methyl groups. On the other hand, the conductivity measurements show a quantitative yield of radical cation in both thiourea and tetramethylthiourea. Taking into account the results of these two independent measurements, we conclude that H atom abstraction at the methyl groups by  $\cdot\text{OH}$  is less than 10%.

**Assignment of the Intermediate  $\lambda_{\text{max}} 400 \text{ nm}$ .** The maximum value of the 400-nm absorbance is dependent on the substrate concentration. When this is reduced from  $10^{-3} \text{ mol dm}^{-3}$  to  $6 \times 10^{-6} \text{ mol dm}^{-3}$ ,  $G\epsilon$  decreases from 4.5 to  $3.2 \text{ cm}^{-1} \text{ kGy}^{-1}$  (values corrected for the loss by bimolecular decay). This reduction cannot be accounted for by the decrease in  $G(OH^\cdot)$  due to the lower solute concentration<sup>23</sup> but rather points to the existence of an equilibrium situation where a dimeric radical cation is formed (reactions 5/–5) which is responsible for the strong absorption observed at 400 nm. A structure with a disulfidic linkage is preferred since formamidinium disulfides are produced in the oxidation of aliphatic thioureas (cf. ref 22). Other canonical structures of the adduct **2** exist, apart from the disulfide radical anionic form, e.g., those that involve the materialization of the radical function on the carbon atom or a form reminiscent of a dimeric sulfide radical cation. The latter form represents the oxidative nature (see below) of **2** particularly well. The equilibrium constant  $K_5 = k_5/k_{-5}$  can be obtained from

(27) Robinson, R. A.; Stokes, R. H. *Electrolyte Solutions*; Butterworth: London, 1959; 571 pp.



**Table 1.** Compilation of Rate Constants Determined in This Study

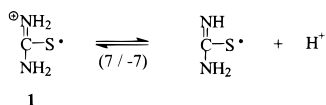
no.	reactions	rate constants
3	thiourea + $\cdot\text{OH} \rightarrow$ thiourea-OH adduct	$1.2 \times 10^{10} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
3'	$\text{Me}_4(\text{thiourea}) + \cdot\text{OH} \rightarrow \text{Me}_4(\text{thiourea})\text{-OH adduct}$	$8.0 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
4	thiourea-OH adduct $\rightarrow$ thiourea $^{+\bullet}$ + OH $^-$	$6 \times 10^7 \text{ s}^{-1}$
5	thiourea $^{+\bullet}$ + thiourea $\rightarrow$ (thiourea) $_2^{+\bullet}$	$5.0 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
-5	(thiourea) $_2^{+\bullet} \rightarrow$ thiourea $^{+\bullet}$ + thiourea	$9.1 \times 10^3 \text{ s}^{-1}$
9	$2(\text{thiourea})_2^{+\bullet} \rightarrow [(\text{thiourea})_2]^{2+} + 2\text{thiourea}$	$9.0 \times 10^8 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
9'	$2(\text{Me}_4(\text{thiourea}))_2^{+\bullet} \rightarrow [(\text{Me}_4(\text{thiourea}))_2]^{2+} + 2\text{Me}_4(\text{thiourea})$	$1.3 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
10	(thiourea) $_2^{+\bullet}$ + OH $^- \rightarrow$ (thiourea) $_2^{\bullet}$	$5.9 \times 10^7 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
12	thiourea $^{+\bullet}$ + OH $^- \rightarrow$ thiyl radical + H $_2\text{O}$	$2.8 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
13	thiyl radical + thiourea $\rightarrow$ neutral dimer radical	$6 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
14	(thiourea) $_2^{+\bullet}$ + O $_2 \rightarrow [(\text{thiourea})_2\text{O}_2]^+$	$1.2 \times 10^7 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
-14	$[(\text{thiourea})_2\text{O}_2]^+ \rightarrow (\text{thiourea})_2^{+\bullet} + \text{O}_2$	$2 \times 10^3 \text{ s}^{-1}$
21	(thiourea) $_2^{+\bullet}$ + O $_2^{\bullet-} \rightarrow 2\text{thiourea} + \text{O}_2$	$4.5 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
21'	$(\text{Me}_4(\text{thiourea}))_2^{+\bullet} + \text{O}_2^{\bullet-} \rightarrow 2\text{Me}_4(\text{thiourea}) + \text{O}_2$	$3.8 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
22	(thiourea) $_2^{+\bullet}$ + N $_3^- \rightarrow 2\text{thiourea} + \text{N}_3^{\bullet}$	$4 \times 10^6 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$
23	$(\text{Me}_4(\text{thiourea}))_2^{+\bullet} + \text{PhO}^- \rightarrow 2\text{Me}_4(\text{thiourea}) + \text{PhO}^{\bullet}$	$3.0 \times 10^8 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$

eq 6, where  $A_0$  is the absorbance at 400 nm in thiourea solution of  $2 \times 10^{-3} \text{ mol dm}^{-3}$  and  $A$  is the absorbance (corrected for  $G(\cdot\text{OH})$  according to ref 23) at a given concentration of thiourea.

$$A_0/A = 1 + K_5^{-1}[\text{thiourea}]^{-1} \quad (6)$$

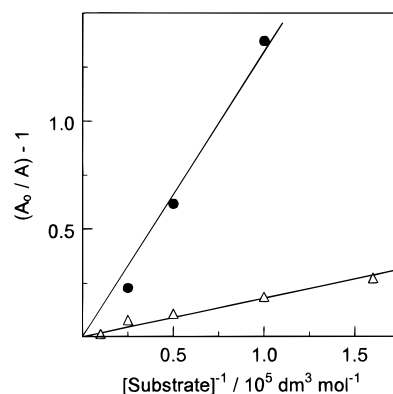
In Figure 4, the term  $(A_0/A) - 1$  is plotted against the reciprocal of the thiourea concentration. From the reciprocal value of the slope of this linear plot,  $K_5 = 5.5 \times 10^5 \text{ dm}^3 \text{ mol}^{-1}$  is obtained for thiourea. The corresponding value of  $K_5 = 7.6 \times 10^4 \text{ dm}^3 \text{ mol}^{-1}$  for tetramethylthiourea (cf. Figure 4) is lower than that of thiourea, indicating that, in this case, the equilibrium reaction 5/-5 is less favorable toward the formation of the secondary, dimeric radical cation.

The fact that, for thiourea, the relationship shown in Figure 4 retains its linear appearance near a concentration of  $10^{-5} \text{ M}$  suggests that the spontaneous deprotonation of the radical cation **1** of thiourea does not significantly contribute to its disappearance. This implies, on the basis of a comparison of the rates of reactions 5 and 7, that the lower limit of the  $\text{p}K_a$  of **1** should be in the vicinity of 6. A  $\text{p}K_a$  value of  $5 \pm 2$  has been estimated on thermochemical grounds.<sup>28</sup>



With the values of  $k_3$  and  $K_5$  obtained above, we have simulated the reaction sequence 3-5 for the absorbance build-up of the dimeric radical cation **2** (cf. Scheme 1, R = H) at 400 nm (cf. Figure 1 inset). The best fit (solid line in the inset of Figure 1) was obtained with  $k_3 = 1.2 \times 10^{10} \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ ,  $k_4 = 6 \times 10^7 \text{ s}^{-1}$ ,  $k_5 = 5.0 \times 10^9 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$ , and  $k_{-5} = 9.1 \times 10^3 \text{ s}^{-1}$ , as well as with the assumption that the absorption coefficient at 400 nm of the monomeric radical cation **1** is about 10% of that of **2**. These results support the above hypothesis that the reaction of the OH radical with the thioureas leads first to an adduct, albeit a short-lived one (reaction 3). On the other hand, a simulation which pretends direct formation of the ion pair, leaving out the OH adduct altogether, shows a build-up kinetics with a much shorter lag phase that does not match the experimental data shown in the Figure 1 inset.

Although such a simulation has to date not been carried out in the case of tetramethylthiourea, a similar mechanism can be assumed since the build-up behavior of the intermediate at 450 nm resembles that from thiourea.

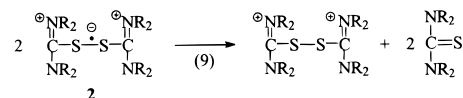


**Figure 4.** Dependence of the term  $(A_0/A) - 1$  at 400 nm ( $\Delta$  for thiourea) and at 450 nm ( $\bullet$  for tetramethylthiourea) on the reciprocal of substrate concentration in the pulse radiolysis of  $\text{N}_2\text{O}$ -saturated solution of thiourea and tetramethylthiourea solution ( $6 \times 10^{-6}$  to  $2 \times 10^{-3} \text{ mol dm}^{-3}$ ) at pH 6.6.

**Bimolecular Decay of the Dimeric Radical Cation.** In the absence of additives that are reactive toward these intermediates, the decay of the absorption near 400 nm exhibits second-order kinetics. The fact that these intermediates carry unit charge of the same sign is corroborated by the behavior of the bimolecular decay rate constants,  $2k$ , which show the characteristic dependence on the ionic strength  $\mu$  (added electrolyte,  $\text{NaClO}_4$ ) and fulfill expression 8<sup>29</sup> (Figure 5).

$$\log(k/k_0) = 1.02z_1z_2\mu^{1/2}/(1 + \mu^{1/2}) \quad (8)$$

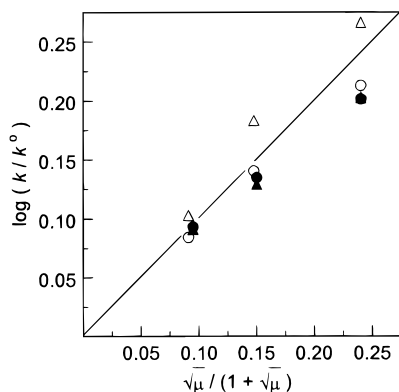
The slope of the plot in Figure 5 is positive and near unity, indicating that the numbers of the electric charge  $z_1$  and  $z_2$  (expression 8) of the reaction partners are unity and of the same sign, in agreement with a bimolecular decay as proposed in reaction 9.



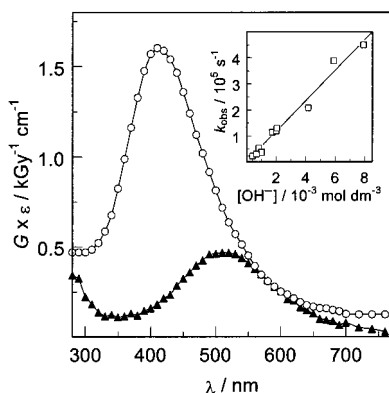
Taking  $G(\text{total radicals}) = 6.4 \times 10^{-7} \text{ mol J}^{-1}$  (this includes the contribution from the primary H atoms), a bimolecular decay rate constant of  $2k_9 = 9.0 \times 10^8 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  ( $1.3 \times 10^9 \text{ dm}^3$

(28) Bordwell, F. G.; Algrim, D. J.; Harrelson, J. A., Jr. *J. Am. Chem. Soc.* **1988**, *110*, 5903-5904.

(29) Lin, S. H.; Li, K. P.; Eyring, H. In *Physical Chemistry. An Advanced Treatise*; Eyring, H., Ed.; Academic Press: New York, 1975; pp 1-58.



**Figure 5.** Ionic strength effect on the rate constant of termination of the radical cation from thiourea monitored at 400 nm (●, ○) and tetramethylthiourea (450 nm, △, ▲), observed at pH 6.6 (open symbols) and at pH 3.1 (filled symbols).



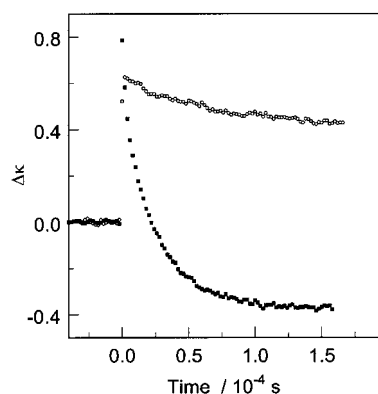
**Figure 6.** Transient absorption spectra obtained in the pulse radiolysis (6 Gy/pulse) of an  $\text{N}_2\text{O}$ -saturated solution of thiourea ( $2 \times 10^{-3} \text{ mol dm}^{-3}$ ) at pH 11.9. Initial spectrum (○), 1  $\mu\text{s}$  after the pulse. Final spectrum (●), 13  $\mu\text{s}$  after the pulse. Inset: dependence of  $k_{\text{obs}}$  of absorption decay at 400 nm on hydroxide concentration.

$\text{mol}^{-1} \text{ s}^{-1}$  for the tetramethyl derivative) is obtained under conditions (neutral, salt-free) where the ionic strength is very low ( $\mu < 10^{-6} \text{ mol dm}^{-3}$ ).

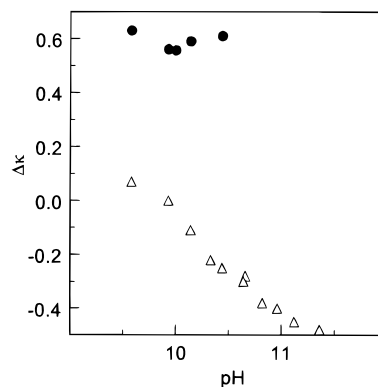
**Transformation of the Radical Cations (Monomer and Dimer) in Basic Solution.** With basic solutions of tetramethylthiourea, the UV spectrum and the conductance signal do not change as the pH is varied (within a range of pH 7–11.2). In contrast, in basic solutions of thiourea above pH 10, the absorption at 400 nm decays rapidly by first-order kinetics within a time where the bimolecular decay of the intermediates is still relatively minor. At the end of this transformation, a new species with  $\lambda_{\text{max}}$  at 510 nm ( $\epsilon = 750 \text{ dm}^3 \text{ mol}^{-1} \text{ cm}^{-1}$ ) is observed (Figure 6). The rate of the absorption decay at 400 nm is proportional to the  $\text{OH}^-$  concentration (Figure 6, inset). From the slope of this plot, the bimolecular rate constant  $5.9 \times 10^7 \text{ dm}^3 \text{ mol}^{-1} \text{ s}^{-1}$  is obtained.

These optical detection results, and the conductivity ones described below for thiourea, are summarized in Scheme 2. The  $\text{OH}^-$ -induced 400-nm absorption decay is ascribed to the reaction of  $\text{OH}^-$  with the secondary, dimeric radical cation **2** to give the neutral species **3** (reaction 10), since under these conditions (high thiourea concentration) the absorption near 400 nm (due to **2**) is already fully developed at the end of the pulse.

Conductance measurements *also* reveal that, in basic solutions, the radical cation of thiourea shows a pronounced difference from that of tetramethylthiourea. In Figure 7, the actual conductance change following an electron pulse in thiourea solution at pH 10.8 is shown to decrease from a positive

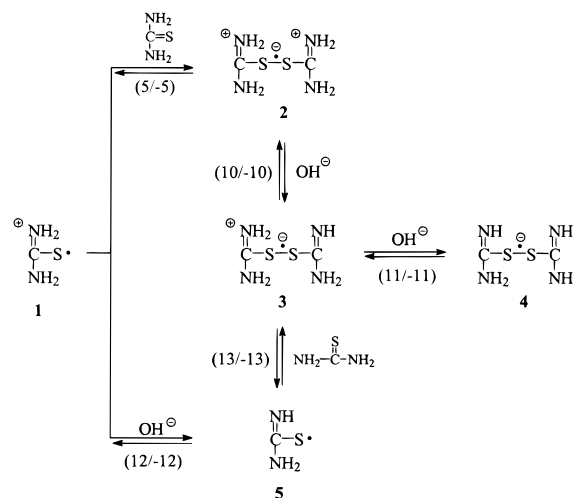


**Figure 7.** Conductance change following an electron pulse of  $\sim 4 \text{ Gy}$  pulse in  $\text{N}_2\text{O}$ -saturated solutions of  $2 \times 10^{-3} \text{ mol dm}^{-3}$  thiourea (■, lower curve) and of  $10^{-3} \text{ mol dm}^{-3}$  tetramethylthiourea (○, upper curve) at pH 10.8.



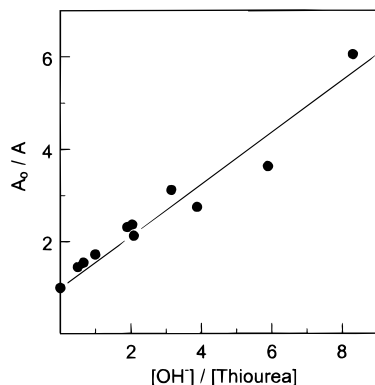
**Figure 8.** Conductance change in  $\text{N}_2\text{O}$ -saturated solution of thiourea ( $2 \times 10^{-3} \text{ mol dm}^{-3}$ ) at 20  $\mu\text{s}$  after the pulse (●) and at the end of the  $\text{OH}^-$ -induced transformation (△, cf. Figure 7) (4 Gy/pulse).

## Scheme 2



value to a negative value in the course of this transformation. For comparison, in tetramethylthiourea solution at identical pH the corresponding change is much smaller and is due to the bimolecular decay only (Figure 7).

In Figure 8, the initial conductance gain and final conductance loss in basic solutions of thiourea are shown over a range of pH values. It appears that, at pH < 10, a neutral species, e.g., the dimeric radical **3**, is formed as the initial conductance gain decays to zero within about 10 ms. At pH > 10, the initial conductance gain decays rapidly and attains a negative value. At pH 11.4, the net loss in conductance reaches a value of



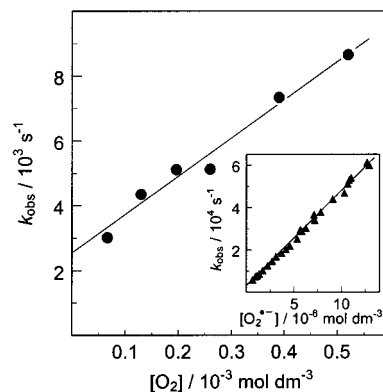
**Figure 9.** Competition plot of the reactions of the radical cation **1** with  $\text{OH}^-$  and with thiourea. The absorbance  $A$  at 400 nm was taken 1  $\mu\text{s}$  after the pulse ( $\sim 2$  Gy/pulse) in  $\text{N}_2\text{O}$ -saturated solution of thiourea ( $2 \times 10^{-4}$ – $2 \times 10^{-3}$  mol  $\text{dm}^{-3}$ ) at pH 10.8–12.  $A_0$  is the absorbance at 400 nm at neutral pH.

$G(-\text{OH}^-) \approx G(\text{OH}^{\bullet})$ ; i.e., the neutral species **3** consumes 1 molar equiv of  $\text{OH}^-$ . The explanation for this is that a further deprotonation, e.g., reaction 11, takes place to give the dimeric radical anion **4**. Obviously, such a deprotonation reaction is not possible in the case of tetramethylthiourea, where only the bimolecular decay is observed (Figure 7). The new species with  $\lambda_{\text{max}}$  at 510 nm observed at the end of  $\text{OH}^-$ -induced transformation (cf. Figure 6) at pH > 10.5 can thus be attributed to the dimeric radical anion **4**.

Parallel to the reaction of  $\text{OH}^-$  with the dimeric radical cation **2** (reaction 10), there is also the reaction of  $\text{OH}^-$  with the monomeric radical cation **1** to give the neutral thiyl radical **5** (reaction 12), which will compete with dimer formation (reaction 5). Indirect evidence for reaction 12 is in the fact that the initial absorbance of **2** at 400 nm decreases with increasing  $\text{OH}^-$  concentration. By measuring the absorbance  $A$  at 400 nm (at 1  $\mu\text{s}$  after the pulse to minimize loss due to reaction 10) in basic solutions of thiourea, we obtained the competition plot shown in Figure 9. Its slope gives the ratio  $k_{12}/k_5 = 0.55$ . Taking the value of  $k_5 = 5.0 \times 10^9$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  obtained above,  $k_{12} = 2.8 \times 10^9$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  is calculated. This value is much higher than that of the equivalent reaction of the dimer **2** (i.e.,  $k_{10} = 5.9 \times 10^7$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$ , see above) and corresponds more closely to what is expected for such reactions.

To elucidate the fate of the neutral thiyl radical **5**, conditions were chosen such that, at sufficiently low thiourea concentration, >99% of the radical cations **1** react with  $\text{OH}^-$  to give **5**, and <1% of **1** react with thiourea to give **2**. Thus, in a  $2 \times 10^{-5}$  mol  $\text{dm}^{-3}$  thiourea solution at pH 11.9, only the build-up of absorption at 510 nm ( $k_{\text{obs}} = 1.3 \times 10^5$   $\text{s}^{-1}$ ), but no absorption at 400 nm is observed. Nevertheless, the resulting absorption at 510 nm is identical to the corresponding absorption obtained in  $4 \times 10^{-3}$  mol  $\text{dm}^{-3}$  thiourea solution (also at pH 11.9), where half of **1** will react to give **4** through the reactions 5, 10, and 11. These results indicate that the thiyl radical **5** subsequently reacts to give **4**, via an addition to thiourea to give the neutral dimeric thiyl radical **3** (reaction 13), followed by the deprotonation reaction 11 to give **4**. The value of  $k_{\text{obs}} = 1.3 \times 10^5$   $\text{s}^{-1}$  for the build-up at 510 nm may be attributed to the pseudo-first-order rate constant of reaction 13, as the deprotonation reaction 12 is not rate-determining at this high pH. This gives  $k_{13} \approx 6 \times 10^9$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$ , similar in order of magnitude to  $k_5$ .

Furthermore, in a  $2 \times 10^{-5}$  mol  $\text{dm}^{-3}$  thiourea solution at pH 10.3, the build-up of absorbance at 510 nm observed at the end of the  $\text{OH}^-$ -induced transformation (30  $\mu\text{s}$  after the pulse,

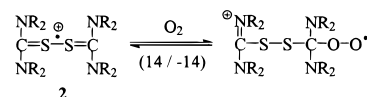


**Figure 10.** Dependence of  $k_{\text{obs}}$  of the disappearance, extrapolated to zero dose, of the thiourea dimeric radical cation, monitored at 400 nm, on the oxygen concentration in  $\text{N}_2\text{O}/\text{O}_2$ -saturated solution of thiourea ( $10^{-3}$  mol  $\text{dm}^{-3}$ ) at pH 6.6. Inset: Pulse radiolysis of  $\text{N}_2\text{O}/\text{O}_2$  (9:1 v/v)-saturated solution of thiourea ( $4 \times 10^{-4}$  mol  $\text{dm}^{-3}$ ) and sodium formate ( $10^{-2}$  mol  $\text{dm}^{-3}$ ) at pH 6.6. Dependence of  $k_{\text{obs}}$  of absorbance decay at 400 nm on superoxide concentration.

$k_{\text{obs}} = 1.4 \times 10^5$   $\text{s}^{-1}$ ) is only  $\sim 60\%$  of the value at pH 11.9. As can be judged from Figure 8, at pH 10.3, the dimeric thiyl radical **3** is only partially deprotonated, and the lower absorbance at 510 nm observed under these conditions lends additional support to our assumption that the neutral species **3** and **5** do not (or only weakly) absorb at this wavelength and that the species with  $\lambda_{\text{max}}$  at 510 nm can be attributed to the dimeric radical anion **4**.

**Reactions of the Dimeric Radical Cations with  $\text{O}_2$ .** In the presence of  $\text{O}_2$ , the rate of decay of the absorption at 400 nm in thiourea solution at neutral pH is slightly faster than that in the absence of oxygen. The rate constant of the reaction of  $\text{O}_2$  with the dimeric radical cation is obtained by extrapolating the observed first-order decay rate constants to zero radical concentration. This is repeated at various  $\text{O}_2$  concentrations. The intercepts so obtained are plotted versus the  $\text{O}_2$  concentration (Figure 10). From the slope of the latter plot, a bimolecular rate constant of  $k_{14} = 1.2 \times 10^7$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  is obtained.

It is assumed that the reaction with  $\text{O}_2$  involves the formation of a peroxy radical (reaction 14). The intercept ( $2 \times 10^3$   $\text{s}^{-1}$ , Figure 10) could signify the existence of a reverse reaction ( $-14$ ) and might be interpreted as  $k_{-14}$  (a reversibility of peroxy radical formation, even at room temperature, has been observed in other cases, cf. refs 30–33). The behavior in the tetramethylthiourea case is similar (not shown).



Radical cations usually do not show reactivity toward  $\text{O}_2$  on the time scale of pulse radiolysis; in the present case, there is reactivity ( $k_{14}$ ), but it is moderate compared to that of other radicals, where rate constants above  $10^9$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  are the norm for the formation of peroxy radicals (cf. ref 34).

(30) Tamba, M.; Simone, G.; Quintiliani, M. *Int. J. Radiat. Biol.* **1986**, *50*, 595–600.

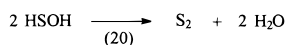
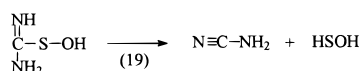
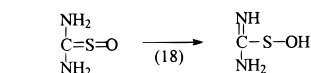
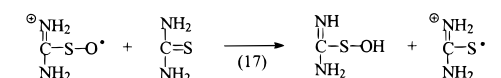
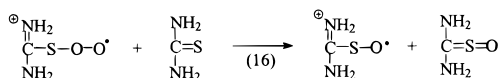
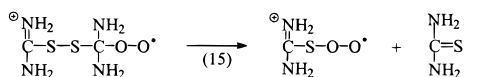
(31) Zhang, X.; Zhang, N.; Schuchmann, H.-P.; von Sonntag, C. *J. Phys. Chem.* **1994**, *98*, 6541–6547.

(32) Pan, X.-M.; von Sonntag, C. *Z. Naturforsch.* **1990**, *45b*, 1337–1340.

(33) Fang, X.; Pan, X.; Rahmann, A.; Schuchmann, H.-P.; von Sonntag, C. *Chem. Eur. J.* **1995**, *1*, 423–429.

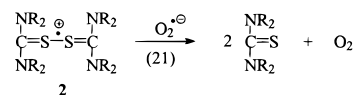
(34) von Sonntag, C.; Schuchmann, H.-P. In *Peroxy Radicals*; Alfassi, Z. B., Ed.; Wiley: Chichester, 1997; pp 173–234.

Puzzlingly, in the presence of oxygen, thiourea may decompose with very high  $G$  values under certain conditions.  $G$  (elemental sulfur) values of up to about  $10^{-3}$  mol  $J^{-1}$  have been observed in oxygenated solutions under conditions of very low dose rate ( $10^{-4}$  Gy  $s^{-1}$ ) and relatively high thiourea concentration (1 M).<sup>8</sup> The mechanism of this chain reaction is unknown. It is conceivable that this chain process is mediated by sulfur peroxy radicals. This kind of radical may be quite reactive (cf. refs 35–37). In the present case, such radicals may be generated via reaction 15. It is tempting to speculate that reactions 16–20 come into play as well.



The formamidinium sulfinyl radical cation may be expected to be an even stronger oxidant than the primary radical cation 1 (reaction 17). Cyanamide is a major product in the  $\text{O}_2$ -assisted oxidative degradation of thiourea.<sup>9,38</sup> There is evidence that hydrogen hydroxysulfide (HSOH) is labile with respect to decomposition (reaction 20) into water and sulfur.<sup>39</sup>

**Reaction with the Superoxide Radical.** Superoxide radicals [ $G(\text{O}_2^{\bullet-}) = 5.6 \times 10^{-7}$  mol  $J^{-1}$ ] and the dimeric radical cation of thiourea **2** [reactions 3–5,  $G(\mathbf{2}) = 7.2 \times 10^{-8}$  mol  $J^{-1}$ ] were generated side by side in  $\text{N}_2\text{O}/\text{O}_2$  (9:1 v/v)-saturated solution of sodium formate ( $10^{-2}$  mol  $\text{dm}^{-3}$ ) and thiourea ( $4 \times 10^{-4}$  mol  $\text{dm}^{-3}$ ) at pH 6.6. The removal of the dimeric radical cation of thiourea **2** by  $\text{O}_2^{\bullet-}$  (reaction 21) was monitored by the decay of its absorption at 400 nm at various  $\text{O}_2^{\bullet-}$  concentrations (dose rate dependent) (Figure 10, inset). From the slope of the linear



plot in the Figure 10 inset, a bimolecular rate constant of  $k_{21} = 4.5 \times 10^9$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  is obtained. As **2** is present at one-eighth of the concentration of  $\text{O}_2^{\bullet-}$ , the contribution of the bimolecular decay of **2** to the slope in the Figure 10 inset is negligible. The intercept at zero  $\text{O}_2^{\bullet-}$  concentration observed in the Figure 10 inset reflects the reaction of **2** with  $\text{O}_2$ . The

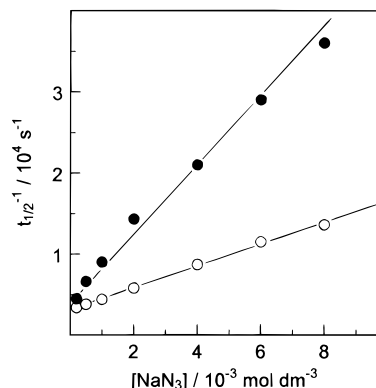
(35) Schuchmann, H.-P.; von Sonntag, C. In *Peroxy Radicals*; Alfassi, Z. B., Ed.; Wiley: Chichester, 1997; pp 439–455.

(36) Asmus, K.-D. In *Active Oxygens, Lipid Peroxides, and Antioxidants*; Yagi, K., Ed.; Japan Sci. Soc. Press: Tokyo/CRC Press: Boca Raton, FL, 1993; pp 57–58.

(37) Schöneich, C. *Antioxid. Health Dis.* **1995**, 2, 21–47.

(38) Schenck, G. O.; Wirth, H. *Naturwissenschaften* **1953**, 40, 141.

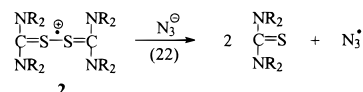
(39) Heinze, E. *J. Prakt. Chem.*, N. F. **1919**, 99, 109–179.



**Figure 11.** Reaction of the thiourea (●) and tetramethylthiourea (○) radical cation with azide, monitored by the disappearance of the absorption at 400 nm (thiourea,  $10^{-3}$  mol  $\text{dm}^{-3}$ ) and at 450 nm (tetramethylthiourea,  $10^{-3}$  mol  $\text{dm}^{-3}$ ) in  $\text{N}_2\text{O}$ -saturated solution at pH 9–10,  $\sim 3$  Gy/pulse.

corresponding value of  $k_{21}$  for tetramethylthiourea is  $3.8 \times 10^9$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$ .

**Reaction with Azide.** The oxidation of thiourea derivatives with the oxidizing radicals  $\text{Cl}_2^{\bullet-}$ ,  $\text{Br}_2^{\bullet-}$ , and  $\text{SO}_4^{\bullet-}$  has been reported, while their oxidation by  $\text{N}_3^{\bullet}$  was not observed, except in the case of the thiourea-related compound thiosemicarbazide,<sup>15,16</sup> even though one would have expected  $\text{N}_3^{\bullet}$  (reduction potential, 1.33 V)<sup>40</sup> to be strong enough to oxidize thiourea; estimates of values for the reduction potentials of some thioureas are reported and fall in the vicinity of 1.1 V.<sup>28</sup> We have exposed thiourea and tetramethylthiourea to the azide radical and also have not observed any effect. This raises the question of the existence of the reverse reaction, i.e., reaction 22.



This is, indeed, observed. From a plot of  $k_{\text{obs}}$  of the absorption decay near 400 nm versus the azide concentration (Figure 11), one determines  $k_{22} = 4 \times 10^6$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  for thiourea and  $10^6$   $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  for tetramethylthiourea. This indicates that the thiourea radical cations are fairly strong oxidants. The intercept in Figure 11 is due to the contribution of termination reactions to the decay. The existence of a reverse reaction would also add to the intercept in Figure 11. However, any sizable contribution by a reverse reaction (–22) can be ruled out since the 400-nm absorption decays to zero (not shown); a reversibility would be reflected in the existence of a nonzero “plateau”.

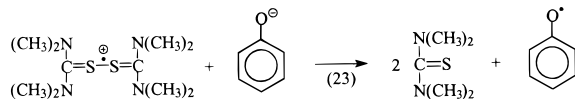
The absence of reactivity of the azide radical (reduction potential 1.33 V)<sup>40</sup> toward thiourea compared to the presence of reactivity<sup>15,16</sup> shown by  $\text{Br}_2^{\bullet-}$  (1.66 V),  $\bullet\text{OH}$  (1.9 V), and  $\text{Cl}_2^{\bullet-}$  (2.3 V) (cf. ref 40) suggests that the reduction potential of the thiourea dimer radical cation is bracketed by 1.33 and 1.66 V. However, in another context, the azide radical has been shown to be unreactive toward dimethyl sulfide with respect to radical adduct formation, even though a species (iodine atom) with a very similar reduction potential does form an adduct;<sup>41</sup> quite often, oxidation reactions proceed via the heterolysis of intermediate adducts and not via direct electron transfer. In fact, the sequence of reactions 3 and 4 presents another example of this case.

(40) Wardman, P. *J. Phys. Chem. Ref. Data* **1989**, 18, 1637–1755.

(41) Merenyi, G.; Lind, J.; Engman, L. *J. Phys. Chem.* **1996**, 100, 8875–8881.



**Oxidation of Phenolate Ion by the (Dimeric) Tetramethylthiourea Radical Cation.** In an  $N_2O$ -saturated solution of tetramethylthiourea ( $10^{-2}$  mol  $dm^{-3}$ ) containing phenol ( $2.2 \times 10^{-4}$ – $10^{-3}$  mol  $dm^{-3}$ ) at pH 11, the absorption spectrum of the dimeric radical cation (cf. reactions 3–5) with  $\lambda_{max}$  at 450 nm observed 1  $\mu s$  after the pulse disappears rapidly. At the end of this process, a new species with  $\lambda_{max}$  at 385 and 402 nm, characteristic of the phenoxyl radical,<sup>42,43</sup> is observed. The  $k_{obs}$  of the absorption decay at 450 nm is proportional to the phenol concentration. From the slope of this linear plot (data not shown), the bimolecular rate constant of the oxidation of phenolate to phenoxyl radical by the radical cation of tetramethylthiourea (reaction 23)  $k_{23} = 3.0 \times 10^8$   $dm^3$   $mol^{-1}$   $s^{-1}$  is obtained.



**Final Products.** There is circumstantial evidence<sup>10</sup> that essentially the only primary product of thiourea radiolysis in aqueous solution is the formamidine disulfide. The formation of elemental sulfur and other products such as guanidylthiourea, guanidine, cyanamide, and dicyandiamide that have been observed in thiourea radiolytic systems in the past<sup>9,10</sup> can in most cases be explained through the rearrangement and subsequent hydrolytic<sup>14</sup> or oxidative degradation of the primarily formed formamidine disulfide.

The formamidine disulfide dication is stable under acidic conditions<sup>14,22</sup> but begins to undergo hydrolysis to the unstable

base if the pH of the solution is above neutral. Similar observations, the production of elemental sulfur in particular, have been reported in the base-catalyzed decomposition of certain substituted formamide disulfides and related compounds.<sup>44</sup> This decomposition proceeds after deprotonation<sup>14</sup> where possible. The decomposition of the homologous compound derived from tetramethylthiourea must take a different course, presumably via hydroxide addition at higher pH.

**General Importance of the Thiourea Radical Cation in Thiourea Oxidation.** The radiolytic oxidation of thioureas to the corresponding formamidine disulfides represents a special case within a whole class of oxidation reactions leading to the same dimer (cf. ref 22 and references therein), among them anodic oxidation,<sup>20</sup> oxidation via nitrosation,<sup>45</sup> where the intermediate S-nitroso cation may be expected to lose nitric oxide by homolysis,<sup>46</sup> and action by oxidants such as chlorine dioxide<sup>47</sup> or methylene blue.<sup>48</sup> In radiolysis, the intermediacy, over a large pH range, of the dimeric thiourea radical cation has now been unequivocally established. It can be fairly assumed that this species plays a dominant role in many of these other processes as well.

**Acknowledgment.** This work is dedicated to Dr. K. U. Ingold on the occasion of his 70th birthday. It has been supported by the European Commission, Project No. F14P-CP95-0011.

JA983275B

(44) Fromm, E. *Liebigs Ann. Chem.* **1906**, 348, 144–160.

(45) Williams, D. L. H. *J. Chem. Soc., Perkin Trans. 2* **1977**, 128–132.

(46) Schulz, U.; McCalla, D. R. *Can. J. Chem.* **1969**, 47, 2021–2027.

(47) Rabai, G.; Wang, R. T.; Kustin, K. *Int. J. Chem. Kinet.* **1993**, 25, 53–62.

(48) Bhagava, R.; Mishra, K. K. *Phosphorus, Sulfur, Silicon* **1991**, 63, 71–79.

(42) Land, E. J.; Porter, G.; Strachan, E. *Trans. Faraday Soc.* **1961**, 57, 1885–1893.

(43) Bansal, K. M.; Fessenden, R. W. *Radiat. Res.* **1976**, 67, 1–8.