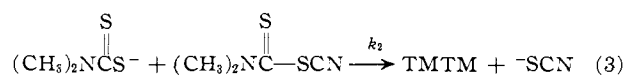
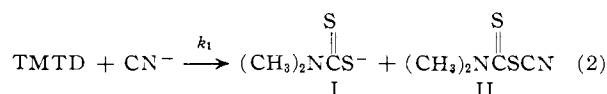


TABLE I
 REACTION OF TETRAMETHYLTHIURAM DISULFIDE WITH POTASSIUM CYANIDE

TMTD, $M \times 10^3$	KCN, $M \times 10^4$	Solution	Salt, $M \times 10^3$	$T, ^\circ C.$	$k_1, M^{-1} \text{ sec.}^{-1} \times 10^{-1}$	$k_2, M^{-1} \text{ sec.}^{-1} \times 10^{-1}$
4.46	12.6	MeOH		25.0	1.5	
3.20	7.5	MeOH	^a	25.0	1.5	
3.03	15.3	MeOH-H ₂ O ^a	50% phosphate ^b	25.0	0.7	
1.00	2.50	MeOH ^c		21.6	1.3	0.7
1.02	5.60	MeOH ^c		21.6	1.2	0.7
2.0	43.6	MeOH		21.6	1.2	
9.98	5.60	MeOH ^c		21.6	1.1	0.6
12.1	10.9	MeOH ^c		21.6	1.2	0.4
2.02	10.0	MeOH	KSCN 6.00	21.6	1.1	
2.02	10.9	MeOH	KSCN 1.20	21.6	1.3	
2.02	10.9	MeOH	KCl 2.50	21.6	1.2	
2.02	10.9	MeOH	KCl 5.0	21.6	1.1	
2.02	10.9	MeOH	KCl 12.0	21.6	1.0	
4.37	5.97	MeOH		15.0	0.83	
4.57	6.54	MeOH-H ₂ O ^a	10% ^a	15.0	.80	
4.95	7.38	MeOH-H ₂ O	40%	15.0	.72	
1.01	5.75	MeOH-H ₂ O	70%	15.0	.70	
2.02	1.09	MeOH		11.4	.60	
2.03	1.08	MeOH		11.4	.63	
2.05	10.0	MeOH ^c		5.0	.47	0.14
2.53	9.80	MeOH-H ₂ O	70%	5.0	.36	
2.40	10.7	MeOH-H ₂ O	10%	5.0	.43	
10.0	1.30	MeOH-H ₂ O ^c	10%	5.0	.40	0.10

^a Volume %; ^b pH 5.6; ^c A small amount of sodium methoxide was added to suppress the hydrolysis of the cyanide and to prevent the decomposition of the dithiocarbamate.



mechanism. An attempt to isolate the thiocarbonyl thiocyanate was made by mixing TMTD and potassium cyanide in methanol followed by the addition of hydrochloric acid in a good hood to destroy the dithiocarbamate¹¹ and remove cyanide ion. An unstable impure fraction was obtained that had an infrared band near 4.5μ which might be the stretching frequency of the carbon-nitrogen triple bond in II. However, the zinc salt of the dithiocarbamate was isolated in the reaction of zinc cyanide with TMTD. In this case the zinc salt was obtained in high yield. Another unstable, impure fraction was obtained that might have been II. It produced thiocyanate ion upon standing in numerous solvents. Thus there is good evidence both spectroscopic and chemical for the two-step processes 2 and 3.

Kinetics.—A detailed analysis of the time behavior of TMTD and cyanide supports the two-step mechanism 2 and 3. Some experiments were run with equivalent concentrations and the data fitted with the equations of the appendix. It was more convenient to run the reaction with a large excess of cyanide. The data are presented in Table I. All kinetic data were processed on a computer. See Fig. 2 for a typical experiment.

The reaction is first order in TMTD over a 12-fold variation in initial concentration and nearly first order in cyanide ion. Control of cyanide ion with low pH means that the dithiocarbamate will decompose.

The rate constant k_1 is depressed by the addition of inert salts and by dilution of the methanol with

(11) (a) G. D. Thorn and R. A. Ludwig, "The Dithiocarbamates and Related Compounds," Elsevier Press, New York, N. Y., 1962; (b) M. Hallaway, *Biochim. Biophys. Acta*, **36**, 538 (1959); (c) P. Zuman and R. Zahradnik, *Z. physik. Chem.*, **208**, 135 (1957); (d) R. Zahradnik and P. Zuman, *Chem. Listy*, **52**, 231 (1958).

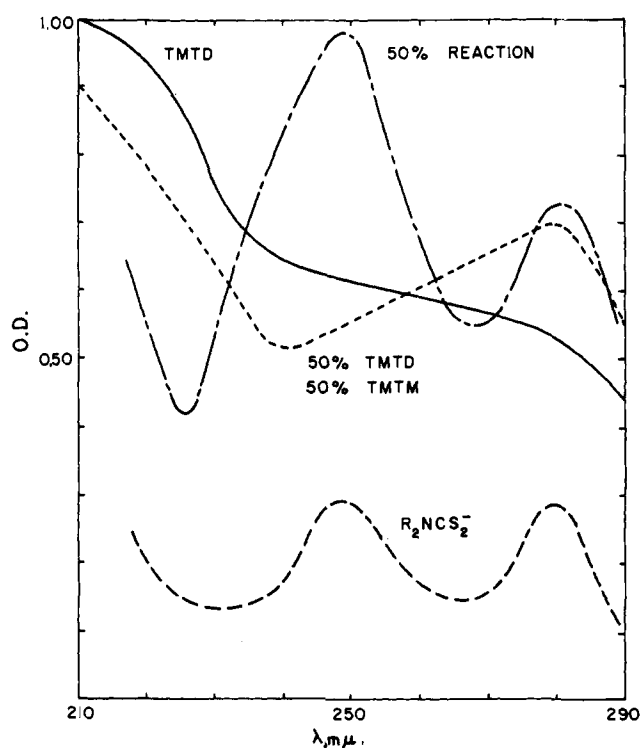
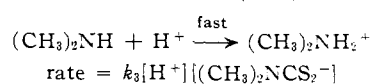
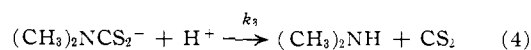


Fig. 1.—Ultraviolet spectra in methanol of TMTD, the synthetic mixture of one part TMTD and one part TMTM, and the observed spectrum of TMTD + KCN at 50% reaction. Also included is the spectrum of dimethyldithiocarbamate. Concentrations are $5.0 \times 10^{-5} M$.

water. Both effects are consistent with the reaction of a charge ion with a neutral molecule with charge



$k_3 = 1.6 \times 10^{+2} M^{-1} \text{ sec.}^{-1}$ at 25° in water (pH range 5.9 to 8.4)

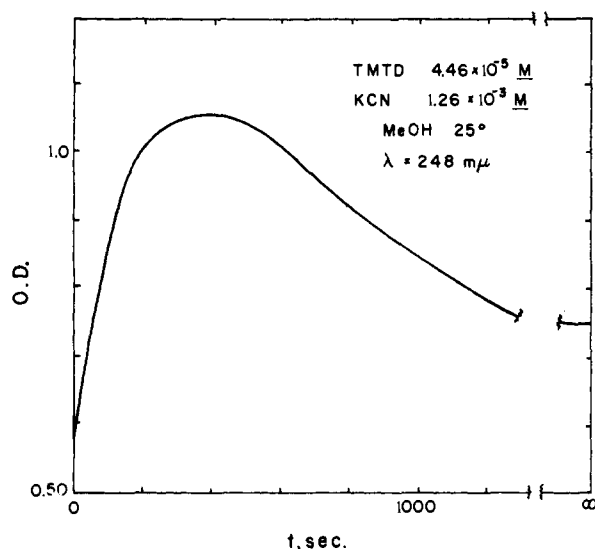


Fig. 2.—Optical density vs. time at 248 m μ for a solution of 4.46×10^{-5} M TMTD and 1.26×10^{-3} M KCN in methanol at 25°.

being dispersed in the activated complex. The activation energy for the reaction is low, being about 10 kcal./mole, a value typical for the reaction of cyanide ion with labile sulfur-sulfur bonds.⁴⁻⁶

Discussion

Thiophiles and Displacement Reactions.—Cyanide ion is easily oxidized and quite basic, therefore it is a very good nucleophile for a displacement reaction on a sulfur-sulfur bond.^{4,5,12} Expressed in the quantitative terms of the oxibase scale,^{4,5} $E = 2.79$ and $H = 10.88$. Yet cyanide ion will readily displace another very good nucleophilic material, the dithiocarbamate.¹³ Using the data of Ogston on the displacement of diethyl-dithiocarbamate on the mustard cation, the value of $E = 2.39$ and $H = 5.74$ can be calculated.⁵ These values will be good estimates for the dimethyl derivative.

Thus in terms of the oxibase scale and the E - and H -values of the nucleophile and the legation,¹⁴ both the E decreases and the H decreases. Both of these

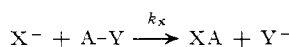
(12) The use of oxidation potentials of $2X^- \rightarrow X_2 + 2e^-$ to correlate relative displacement orders was first made by O. Foss (*Acta Chem. Scand.*, **1**, 307 (1947); in "Organic Sulfur Compounds," N. Kharasch, Ed., Vol. I, 1961, pp. 75-96). Edwards (J. O. Edwards, *J. Am. Chem. Soc.*, **76**, 1540 (1954); **78**, 1819 (1956); J. O. Edwards and R. G. Pearson, *ibid.*, **84**, 16 (1962)) then suggested his empirical double basicity scale combining both the oxidation potential and the basicity. We call the scale the oxibase scale.

A theoretical justification using nonequilibrium thermodynamics has been made.^{4,5} We would rather use the electrode potentials than the polarizability term of Edwards.

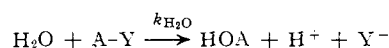
The linear-free energy relation is

$$\log(k_x/k_{H_2O}) = \alpha E + \beta H$$

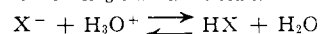
where k_x is the second-order rate constant of



and k_{H_2O} refers to



The E -value is the oxidative electrode potential relative to water and the H is the pK_a of X^- plus the log of water concentration. Thus H is the pK_a of a single X^- compared to a single water molecule.

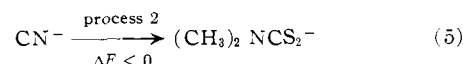


The constants α and β are characteristic of the substrate $A-Y$. A novel interpretation^{4,5} has been made concerning α and β . The term α is a measure of the ease of reduction of $A-Y$ and is proportional to the E -value of the leaving group. The term β is a measure of the acidity and the amount and degree of charge near the reaction center in $A-Y$.

(13) A. G. Ogston, E. R. Holiday, J. St. L. Philpot, and L. A. Stocken, *Trans. Faraday Soc.*, **44**, 45 (1948).

(14) A word suggested in ref. 4 and δ for a leaving group anion, from *lego* (Greek), to leave off. A neutral species would be a *legite*.

changes result in a favorable free energy change and process is spontaneous.

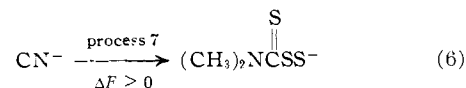


$$E \ 2.79 \longrightarrow 2.39 \text{ decrease}$$

$$H \ 10.88 \longrightarrow 5.74 \text{ decrease}$$

It is quite obvious from the use of the oxibase scale that cyanide would not add to the C=S bond and

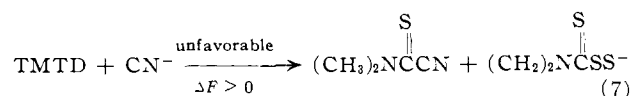
displace the $(CH_3)_2NCSS^-$ anion. An estimate of



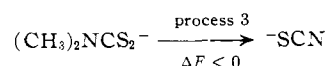
$$E = 2.79 \longrightarrow 2.9 \text{ increase}$$

$$H = 10.88 \longrightarrow 11 \text{ to } 13 \text{ increase}$$

the oxidative dimerization potential of this polysulfide^{4,5} ion would place E near 2.9. An estimate of the pK_a would place H between 11 and 13. Thus the E increases and the H probably increases. As most of the effect^{4,5} is in the change of E -values, this process 7 has an unfavorable ΔF and would not occur.



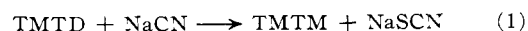
The thiocarbamate will easily displace the thiocyanate legation. In this case the changes of both E and H are quite favorable.



$$E \ 2.39 \longrightarrow 1.83 \text{ decrease}$$

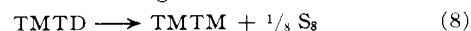
$$H \ 5.74 \longrightarrow 1 \text{ decrease}$$

The Energetics of the System.—In 1960, Davis,¹⁰ discussing the thermochromism properties of TMTD, reported the heat of reaction 1 in methanol using a



$$\Delta H = -19.8 \pm 0.4 \text{ kcal./mole}$$

simple calorimeter. Using this datum and the heats of formation of sodium thiocyanate and sodium cyanide, an estimate¹⁰ was given for the ΔH of



$$\Delta H = -2.4 \text{ kcal./mole}$$

Recently McCullough¹⁵ measured the heats of combustion and of formation of TMTD and TMTM using a precision rotating bomb calorimeter. Using their values of ΔH_f° (solid) of 11.44 ± 0.30 kcal./mole for TMTM and 9.74 ± 0.41 kcal./mole for TMTD, we have calculated the ΔH of (1) to be -17.20 ± 0.70 and the ΔH of (8) to be -1.70 ± 0.71 kcal./mole. The agreement of our data, simply obtained, with more exacting data is quite good. Some correction would be needed as the ΔH_f° 's refer to the solids and our values refer to homogeneous solution in methanol.

An Estimate of the Sulfur-Sulfur Bond Length by Kinetic Methods.—The activation energy for the initial reaction 2 is low. Using the empirical correlation¹ between the activation energy and the sulfur-sulfur bond length, an estimate of 2.12 ± 0.04 Å. is placed on the sulfur-sulfur bond length in TMTD.¹⁶ Such a value is consistent with the general reactivity and thermal

(15) W. D. Good, J. L. Lacina, and J. P. McCullough, *J. Phys. Chem.*, **65**, 860 (1961).

(16) An X-ray analysis of TMTD has been suggested to Prof. Olav Foss, personal communication, spring, 1963.

instability of TMTD,^{10,17-20} and the resonance stability of the dithiocarbamyl free radical, R₂NCS₂.

Experimental

Materials.—Eastman Kodak Yellow Label tetramethylthiuram disulfide was recrystallized.¹⁰ The monosulfide was isolated from the reaction of TMTD (Yellow Label, not purified) with potassium cyanide in methanol. The product was recrystallized from aqueous methanol, isolated in 88% yield (m.p. 105–106°, corrected). Methanol was refluxed over magnesium turning and distilled through an 8-ft. packed column. The fraction boiling at 64.5° (750 mm.) was stored in the dark. Potassium cyanide (Baker reagent) was standardized against silver nitrate. The cyanide solutions were prepared daily. Potassium chloride and thiocyanate (Baker reagents) were dried at 110°.

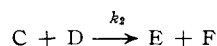
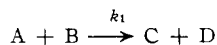
Kinetics.—A Cary Model 14 spectrometer or a thermostated Beckman DU with 1-cm. cells were used. All kinetic data were treated on a Royal McBee RPC-4000 digital computer using a program written by A. C. An estimate of the standard deviation was also provided.

Isolation.—Zinc cyanide (0.10 mole) was suspended in 600 ml. of methanol. The solvent was then allowed to reflux slowly through the thimble of a Soxhlet extractor containing 0.10 mole of TMTD. After the TMTD had been added, the solution was partially concentrated. The white solid removed by filtration was washed with acetone. From the acetone solution 0.09 mole of the zinc dithiocarbamate was obtained, identical in all respects with a known sample in the infrared, n.m.r., ultraviolet, and melting point.

Further concentration of the original methanol solution resulted in an unstable oil having strong absorption near 4.5 μ. Numerous attempts failed to win a pure product.

Appendix

The differential equations for the kinetic system



have not been solved in closed form. If one uses equivalent concentrations, then A = B, C = D, E = F. In this case



and solution is possible. The solution of system (i) and (ii) has not been reported in the literature.

From i

$$A = \frac{A_0}{1 + A_0 k_1 t} = \frac{A_0}{\tau} \quad (iii)$$

Then

$$\frac{dC}{dt} = k_1 \frac{(A_0)^2}{\tau} - k_2 C^2 \quad (iv)$$

Let

$$C = \frac{1}{k_2 u} \frac{du}{dt} \quad (v)$$

and

$$e^s = \tau \quad (vi)$$

and

$$k_2/k_1 = k \quad (vii)$$

Then iv becomes

$$\frac{d^2 u}{ds^2} - \frac{du}{ds} - ku = 0 \quad (viii)$$

Solution of viii using the differential operator $D = d/ds$ gives

$$u = g_1 \exp \left\{ \frac{1}{2} + \frac{1}{2}(\sqrt{1+4k})s \right\} + g_2 \exp \left\{ \frac{1}{2} - \frac{1}{2}(\sqrt{1+4k})s \right\} \quad (ix)$$

with g_1 and g_2 constants. Defining

$$\lambda = \sqrt{1+4k} \quad (x)$$

and evaluation of g_1 and g_2 with the boundary conditions $\tau = 1$ and $C = 0$ at $t = 0$

$$C = \frac{A_0}{2\tau} k \left[\lambda + 1 - \frac{2\lambda}{1 + \left(\frac{\lambda-1}{\lambda+1}\right)^\lambda} \right] \quad (xi)$$

Then as

$$\begin{aligned} dE &= k_2 C^2 dt = k_1 A^2 dt - dC \\ E &= A_0 \left[1 - \frac{1}{\tau} \right] - C \end{aligned} \quad (xii)$$

The value of k_1 can best be obtained from expression iii obtained from a time study of the concentration of A . The value of k_2 can be estimated by plotting the concentrations of A , C , and E vs. τ^{-1} ; C reaches a maximum at $t_{c \max}$ when dC/dt and $d^2C/dt^2 = 0$.

An analytical expression then relates k_1 , t_{\max} , and k_2 . These operations are quite rapid on a computer.

Acknowledgment.—We wish to thank the Walter Reed Army Institute of Research, the National Science Foundation, the Research Corporation, and the National Institutes of Health for grants and aids providing the equipment and support for these projects. Some experiments were performed by Mr. Harry Rubins and Mr. Steve Olin. Mr. Nakshbendi obtained the kinetic data reported in eq. 4.

(17) J. C. D. Brand and J. R. Davidson, *J. Chem. Soc.*, 15 (1956).

(18) C. Walling, "Free Radicals in Solutions," John Wiley and Sons, Inc., New York, N. Y., 1957, pp. 335–341.

(19) T. E. Ferington and A. V. Tobolsky, *J. Am. Chem. Soc.*, **77**, 4510 (1955); **80**, 3215 (1958).

(20) D. Craig, *Rubber Chem. Tech.*, **30**, 1291 (1957).