A Study of d-Orbital Effects in Esters of Trifluorothiolacetic Acid. A Comparison of Evidence from Dipole Moment and Kinetic Data

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The dipole moments of a series of esters of trifluoroacetic acid and trifluorothiolacetic acid have been determined in benzene solution at 25 $^{\circ}$ C. The results are consistent with conjugative and hyperconjugative electron release by the hydrocarbon groups to the d-orbitals of sulfur. This conclusion is supported by data for the hydrolysis of these compounds in water or aqueous acetone.

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On a déterminé les moments dipolaires d'une série d'esters des acides trifluoroacétique et trifluorothiolacétique dans des solutions benzéniques à 25 °C. Les résultats sont en accord avec le concept que les groupes hydrocarbonés repoussent vers les orbitales d du soufre des électrons par conjugaison et hyperconjugaison. Cette conclusion est confirmée par les données sur l'hydrolyse de ces composés dans l'eau et dans des solutions aqueuses d'acétone.

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Introduction

The dipole moments of thiolesters and similar compounds are usually greater than those of their oxygen analogs and the differences are particularly large for aryl compounds. We have recently demonstrated that this also applies to chloroformate (1; X = Cl) and chlorothiolformate esters (2; X = Cl) (1) and also that the differences decrease as the hydrocarbon groups change from phenyl to secondary alkyl. For this reason, we have proposed that the observed differences between sulfur and oxygen compounds mainly reflect conjugative or hyperconjugative electron release to the d-orbitals of sulfur in the thio compounds. Under appropriate conditions such effects might manifest themselves in kinetic phenomena. Unfortunately, we have found that chloroformates react bimolecularly with water (mechanism 1) (2) whereas the corresponding

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chlorothiolformates react unimolecularly (mechanism 2) (3, 4). Consequently, a comparison of the reactivities of these compounds towards water would hardly constitute a test of d-orbital effects in the thiolesters. However, esters of trifluoroacetic acid and trifluorothiolacetic acid react bimolecularly with water (mechanisms 4 and 5) (5-7) at convenient rates for systematic study. It therefore seemed that a comparison of the evidence obtainable from dipole moment and kinetic data would be a useful way in which to test the importance of d-orbital effects in sulfur compounds. Moreover, since dipole moments have not been previously reported for any of these compounds, their determination seemed to be a worthwhile piece of work in its own right. We have therefore measured these quantities for seven esters of trifluoroacetic acid $(1; X = CF_3)$ and the corresponding derivatives of trifluorothiolacetic acid (2; $X = CF_3$) in dilute benzene solutions. The hydrocarbon groups were varied so as to provide a wide variation in their abilities to donate electrons by conjugation or hyperconjugation to the d-orbitals of sulfur. The rates of hydrolysis of some of the thiol esters and phenyl trifluoroacetate in water or aqueous acetone

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	Y	Boiling point (°C)		Calculated (%)			Found (%)					
R		This work	Literature	Reference	C	Н	F	S	C	Н	F	S
CH ₃	0	44	43	16								
C_2H_5	0	62	61.5	16								
$-C_3H_7$	0	82.5	82	16								
$-C_3H_7$	0	73.5	73	17								
—C₄H ₉	0	104	104	16								
—C₄H ₉	0	87	85-86	18								
C ₆ H ₅	0	66/38 mm	67/40 mm	17								
CH	S	71-72			25.0	2.1	39.6	22.2	25.0	2.1	39.0	21.8
C ₂ H ₅	S	90.5	90.5	. 19								
$-C_{3}H_{7}$	S	112			34.9	4.1	33.1	18.6	34.7	4.0	32.8	18.9
$-C_{3}H_{7}$	S	104			34.9	4.1	33.1	18.6	34.7	4.3	33.4	18.5
—C₄H₄	S	135										
-C ₄ H ₉	S	113			38.7	4.8	30.6	17.2	38.4	4.6	31.0	17.0
C_6H_5	S	109/70 mm	140-150	20								
*General for	rmula, R—Y	(CO)—CF ₃ .										
			TABLE 3. D	Dipole moments	for halo-est	ers in ben	zene solutic	on at 25 °C*	k			
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TABLE 2	Boiling points of	esters of	trifluoroacetic ac	cid and	trifluorothiolacetic acid*

TABLE 3.	Dipole moments	for halo-esters in	benzene solution at 25 °C*
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	μ (D)		μ ()		
R	RO-(CO)-Cl	RS-(CO)-Cl	$\Delta \mu(S-O)^{\dagger}$	RO—(CO)—CF ₃ §	RS—(CO)—CF ₃ §	Δμ(S—O)‡
C _c H _c	2.39	2.86	0.47	2.68	2.95	0.27
CH ₂	2.38	2.69	0.31	2.71	2.84	0.13
C ₂ H _e	2.66	2.84	0.18	3.00	3.07	0.07
C ₂ H ₇	2.70	2.88	0.18	3.02	3.10	0.08
$n - C_4 H_0$	2.71	2.93	0.21	3.10	3.16	0.06
$i - C_2 H_2$	2.86	2.96	0.10	3.08	3.15	0.07
$t - C_4 H_9$		3.20		3.30	3.37	0.07

*Given in Debye unit D. $\uparrow \Delta \mu(S-O) = \mu(RS-CO-CI) - \mu(RO-CO-CI),$ $\downarrow \Delta \mu(S-O) = \mu(RS-CO-CF_3) - \mu(RO-CO-CF_3).$ §This work. [Reference 1.

have also been measured, corresponding data for the oxygen analogs being already available in the literature (5).



$$\begin{bmatrix} 2 \end{bmatrix} RS - C - CI \iff RS - C \equiv 0^+ + CI^- \xrightarrow{H_2O} 0^+ \\ \parallel \\ O \end{bmatrix}$$

 $RSH + CO_2 + HCl$

Experimental

Solvents

The methods used for purifying the benzene, acetone, and water used as solvents in the present studies have been previously described (1–4). Some physical properties of benzene solutions are listed in Table I lodged with the Depository of Unpublished Data.⁶

Materials

The esters were prepared by reacting the corresponding alcohols and thiols with trifluoroacetic anhydride and were purified by distillation before use. Physical properties are recorded in Table 2.

Procedures

These have been previously described (1-4).

Products of Hydrolysis

Experiments were carried out with methyl and isopropyl trifluorothiolacetates as previously described for the corresponding chlorothiolformate esters (3). The thiols produced were characterized by conversion to the corresponding 2,4-dinitrophenyl alkyl sulfides which were then shown to be identical with authentic samples.

Results

Table 1 summarizes the experimental data from which the solution dipole moments (Table 3) have been calculated by Smith's method (8) using eq. 3

[3]
$$\mu = \frac{27kT}{4N} \frac{M_2}{d_1(\varepsilon_1 - 2)^2} (a - v)$$

The symbols have the same meanings as those used in our earlier papers (1). In order to facilitate comparisons, Table 3 also includes previously reported values of the dipole moments of the corresponding chloroformate (9) and chlorothiolformate esters (1).

Table 4 lists rate constants for the hydrolysis of four esters of trifluorothiolacetic acid (2; $X = CF_3$, $R = C_6H_5$, C_2H_5 , $i - C_3H_7$) and the corresponding esters of trifluoroacetic acid (1; $X = CF_3$), the solvent being water in the cases of the alkyl esters and 70% aqueous acetone for the phenyl compounds which react very rapidly with water. The values for the alkyl trifluoroacetates have been calculated from the data of Winter and Scott (5). These workers have shown that *t*-butyl trifluoroacetate hydrolyses, at least partly, by an S_N1 mechanism and we have therefore excluded tertiary esters from our studies. In the case of isopropyl trifluorothiolacetate, complete hydrolysis of 0.002 mol of the ester in 100 ml of water gave only 94.7% of the calculated amount of acid. Nevertheless, the compound appeared to be quite pure since only a single peak was produced when it was examined by vapor phase chromatography. Isopropyl mercaptan was identified as a product of hydrolysis and there was no evidence that isopropyl alcohol was formed, indicating that the possibility of alkyl-sulfur bond fission may be ignored. In individual runs at concentrations of approximately 0.001 M, good first order kinetics were obtained but agreement between the values obtained in different experiments were rather poor. However, by working at lower concentrations of the ester (approximately 0.00005 M), satisfactory agreement between separate runs was achieved. These problems were probably associated with supersaturation in the more concentrated solutions. Despite these difficulties, it is considered that the rate constants for the hydrolysis of isopropyl trifluorothiolacetate are sufficiently reliable for the purposes of the present studies. No similar problems arose with the other compounds which gave excellent, reproducible rate constants and at least 99% of the theoretical amount of acid on hydrolysis.

Discussion

Dipole Moments

The results in Table 3 show three features to which we have previously drawn attention (1).

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⁶Complete set of data may be obtained, at a nominal charge, from the Depository of Unpublished Data, National Science Library, National Research Council of Canada, Ottawa, Canada K1A 0S2.

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Solvent	R	RO—(CO)—CF ₃	RS—(CO)—CF ₃ *	$k_{\rm obs}{}^{\rm 0}/k_{\rm obs}{}^{\rm s}$	
70% aqueous acetone	C ₆ H ₅	82.54*	15.90	0.1926	
Water	CH_3	25.72†	19.55	0.7601	
Water	C_2H_5	9.642†	14.08	1.460	
Water	$i - C_3 H_7$	2.702†	9.753	3.610	

TABLE 4. Rates of hydrolysis of esters of trifluoroacetic acid and trifluorothiolacetic acid in water or 70% aqueous acetone at 6.09 °C

*Values obtained in this work. †Reference 5.

First of all, the dipole moments of the sulfur compounds are larger than those of their oxygen analogs. Then, the monotonous decrease in the values for the oxygen compounds, as the inductive effect of the group R decreases $(t - C_4 H_9 >$ $i-C_{3}H_{7} > C_{2}H_{5} > CH_{3} > C_{6}H_{5}$), is not exactly mirrored in the thiol esters. In particular, the dipole moments of the phenyl thiolesters are greater than those of the methyl thiolesters, a reversal of the situations for the oxygen compounds. Thirdly, the values of $\Delta \mu(S-O)$ decrease as the ability of the group R to donate electrons conjugatively or hyperconjugatively falls off. It is admitted that the trend is less marked for the present fluorinated esters than for the chlorothiolformates, an apparently constant difference of about 0.07 D being reached at the ethyl compounds. However, this small difference may be less than the combined experiment errors⁷ in the quantities subtracted and may not be really constant. Since we have previously presented our interpretation of these trends in some detail (1) it seems unnecessary to repeat them at this time. However, on the basis of earlier conclusions, the results are consistent with the proposal that conjugative and hyperconjugative electron release to the d-orbitals of sulfur are the main reasons for the larger dipole moments of the sulfur compounds. Other explanations are, of course, quite possible. For example, the different sizes of oxygen and sulfur

atoms and different steric effects associated with the hydrocarbon groups (R) could markedly alter bond angles and spacial orientations of the "lone-pair" electrons. However, in this case it would be reasonable to expect that the largest groups ($R = C_6H_5$ and $t-C_4H_9$) would generate the most closely related values of $\Delta\mu(S-O)$, which is not the case. Moreover, the relationships $\mu(C_6H_5O-CO-CF_3) < \mu(CH_3O-CO-CF_3)$ and $\mu(C_6H_5S-CO-CF_3) > \mu(CH_3O-CO-CF_3)$ would still remain unexplained. In the absence of a knowledge of such differences, further speculation would seem to be pointless and a more profitable approach is to consider evidence of a different kind.

Rates of Hydrolysis

Γ

Esters of trifluoroacetic acid react with water at low acid concentrations by the uncatalyzed process shown in [4]. The observed pseudo firstorder rate constant (k_{obs}°) is given by eq. 5.

4] RO-C-CF₃ + H₂O
$$\xrightarrow{k_1}_{k_2}$$

O
O
O
RO-C-CF₃ $\xrightarrow{k_3}$ ROH + CF₃COOH
O
4
5] $k_{obs}^{\circ} = \frac{k_1}{1 + k_2/k_3}$

Esters of trifluorothiolacetic acid react by a slightly different mechanism [6] in which step 2 is subject to acid catalysis. The rate constant is given by eq. 7 which, at the low acid concentrations generated in the present studies (less than $10^{-3} M$) reduces to eq. 8

⁷The calculation of the standard errors in the dipole moments from the measured quantities is a very complicated procedure. However, one of us has previously demonstrated (10) that the dipole moments calculated from measurements made with our equipment are probably good to ± 0.1 D on an absolute scale and to better than this on a relative scale since standardized conditions and procedures were used for all the experiments.

[6] RS-C-CF₃ + H₂O
$$\xrightarrow{k_1'}_{k_2'[H^+]}$$

OH
RS-C-CF $\xrightarrow{k_3'}$ RSH + CF₃COOH
O⁻
5

[7]
$$k_{obs}^{s} = \frac{k_1'}{1 + (k_2'/k_3')H^+}$$

[8] $k_{obs}^{s} = k_{1}' \text{ at low } [\mathrm{H}^{+}]$

These differences have been discussed in detail by Johnson (6).

Esters, including the present trifluoroacetates (1 and 2, $X = CF_3$), and similar compounds are stabilized by conjugation (3; Y = O or S) involving the carbonyl group and the "lone-pair" electrons on the hetero-atom (Y). Similar conjugation is not possible for the tetrahedral intermediates (4 and 5) formed in step 1 of the two reactions. It would therefore be expected that the rates of these steps would largely reflect the differences in initial state conjugation for the oxygen and sulfur compounds. Since sulfur shows less tendency than oxygen to use its "lone-pair" electrons for π -bond formation (11), it follows that the oxygen esters should be more stabilized than the thio-analogs with $k_1 < k_1'$. Equation 5 shows that k_1 is the upper limit for k_{obs}° , the rate constant for hydrolysis of the oxygen esters, while eq. 8 gives $k_1' = k_{obs}^{s}$. It would therefore be expected that the oxygen esters would hydrolyze more slowly than the thiolesters. On the other hand, the replacement of oxygen by sulfur should alter the inductive effects of the substituent groups (R) by roughly similar amounts, if at all, so that the relative orders of reactivities within each series should be rather similar. While the first of these expectations is realized for the ethyl and isopropyl compounds (Table 4), the second one is not realized at all. Moreover, methyl and phenyl trifluoroacetates hydrolyze more rapidly than the corresponding trifluorothiolacetates, the difference being particularly large for the aromatic compounds. The departures from expectation are particularly well shown by a comparison of the changes of relative reactivities $(k_{obs}^{s}/k_{obs}^{o})$ as the substituent (R) changes from phenyl to isopropyl. Clearly, some effect or effects other than those described above must be operating in at least one of these two series of compounds. Moreover, these effects must decrease as the group R changes from phenyl to isopropyl. But, this is just the trend that would be expected if conjugative or hyperconjugative electron release to the d-orbitals of sulfur occurs in the initial states for the hydrolysis of the present compounds. This is reasonable, since the partial positive charge on sulfur (resonance form **3**; Y = S) should facilitate electron release towards this center.

A referee has suggested that for the phenyl compounds the results might be explained in terms of electron withdrawal due to π overlap, which would be more important for oxygen than sulfur. The phenyl trifluorothiolacetate then reacts more slowly than the oxygen analog due to decreased initial-state stability of the latter compound, rather than increased stability of the sulfur compound due to d-orbital participation as we have suggested. However, this explanation cannot apply to the methyl compounds, where the R group is not only electron donating but cannot participate in electron withdrawing π overlap, despite which the oxygen ester is the more reactive. Moreover, this explanation ignores the dipole moment data which indicates that charge separation is greater in C_6H_5 SCOCF₃ than in CH₃SCOCF₃, despite the fact that the phenyl group should withdraw electrons inductively.

No doubt other explanations involving transition state differences could be advanced. However, it is significant that the kinetic data do support the conclusions drawn from comparisons of dipole moments, which must reflect groundstate properties. While previous studies by Bordwell *et al.* (12) have failed to demonstrate that d-orbital effects influence chemical reactivity, Tarbell and co-workers (13) have previously argued that such electron donation to sulfur could account for the inversion of the ratio k_{obs}^{s}/k_{obs}^{o} in the basic hydrolysis of alkyl acetates and thiolacetates.

Lee (14, 15) has recently carried out INDO and CNDO/2 calculations for methyl chloroformate (1; $R = CH_3$, X = Cl) and its various thioanalogs and compared the results with our experimental dipole moment and kinetic data. He concludes that d-orbital effects account for at least some of the differences between the oxygen and sulfur compounds. We are presently carrying out similar calculations for the trifluoroacetates.

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