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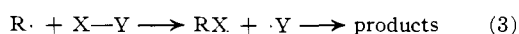
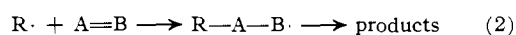
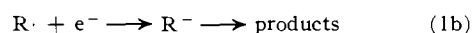
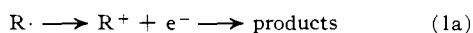
Stable Free Radicals. III. Rates of Reaction of 1-Ethyl-4-carbomethoxy-pyridinyl with Halocarbons

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The rate of halogen abstraction by the stable free radical, 1-ethyl-4-carbomethoxy-pyridinyl, has been measured for the halocarbons, CH₂Cl₂, CH₂Br₂, Br-CH₂Cl, ICH₂Cl, CH₃I, CHCl₃, CDCl₃, and CCl₄. The rates are extremely sensitive to the nature of the halocarbon, a factor of 10⁸ separating CH₂Cl₂ and CCl₄. The sensitivity of the rate to the nature of the halogen, the yield of pyridinium halide (ca. 50%), the small secondary isotope effect ($k_H/k_D = 1.21$ for CHCl₃), the small solvent polarity effect on the rate, and the formation of dihydropyridines support a three-step mechanism: (1) Py· + XC → PyX (covalent) + ·C; (2) Py· + ·C → PyC; and (3) PyX → Py⁺X⁻. The negligible solvent effect on the reaction of the radical with CH₂Br₂ over a Z-value range of 15 demonstrates that the transition state for bromine atom transfer has the same charge separation as the initial state. The contrast between this conclusion and that expected for a contribution of "pyridinium ion stabilization" to the transition state led to the re-evaluation of polar factors in some radical reactions as being attributable to ground state (initial state) effects, as suggested by Semenov, as well as transition state effects. The kinetic data and some simple assumptions lead to maximum values for the stabilization of the product radicals relative to the methyl radical: ·CH₂Cl, 6 kcal./mole; ·CHCl₂, 10 kcal./mole; and ·CCl₃, 17 kcal./mole.

Radical reactions fall into three general classes: electron transfer (eq. 1a and 1b), addition (eq. 2), and atom transfer (eq. 3).⁴ No mechanistic detail is



implied by these equations, each of which could proceed through an intermediate complex containing the reactants or other species. This paper concerns the direct measurement of the atom transfer reaction (eq. 3) in which the radical is 1-ethyl-4-carbomethoxy-pyridinyl and the transferred atom a halogen.

The only previous direct measurements of halogen abstraction reactions were those originally studied by Polanyi and co-workers^{5,6} and later by Warhurst and his collaborators,^{7,8} among others, involving sodium vapor and organic halocarbons. The experimental technique was not amenable to the determination of accurate rate constants, especially for very reactive compounds.⁹ Estimates for the rate constants for the abstraction of bromine from bromotrichloromethane by the 2-trichloromethylcyclohexyl radical (64 l. mole⁻¹ sec.⁻¹) and by the 1-acetoxyethyl radical (2740 l. mole⁻¹ sec.⁻¹) have been obtained by the rotating sector technique,¹⁰ but these are of doubtful precision. Relative rates of halogen abstraction can be measured by competition methods, either by examination of product ratios¹¹ or polymer molecular weight (as in chain transfer).¹²

(1) Alfred P. Sloan Fellow 1960-1964.

(2) The previous papers in this series are: (a) E. M. Kosower and E. J. Poziomek, *J. Am. Chem. Soc.*, **86**, 5515 (1964), and (b) E. M. Kosower and J. L. Cotter, *ibid.*, **86**, 5524 (1964).

(3) Support of the National Science Foundation under Grant GP-251 is gratefully acknowledged.

(4) Although addition could be regarded as an atom transfer reaction in which the transferred atom remains bonded to the product radical, and thus make comparison with one-electron oxidation-reduction reactions facile, it seems more useful to treat the reactions of eq. 2 and eq. 3 separately.

(5) H. von Hartel and M. Polanyi, *Z. physik. Chem.*, **11B**, 97 (1930).

(6) H. von Hartel, N. Meer, and M. Polanyi, *ibid.*, **19B**, 139 (1932).

(7) J. N. Haresnape, J. M. Stevels, and E. Warhurst, *Trans. Faraday Soc.*, **36**, 465 (1940).

(8) E. Warhurst, *Quart. Rev. (London)*, **5**, 44 (1951).

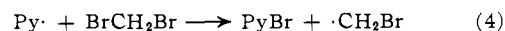
(9) Detailed information about such reactions can be obtained by using crossed monoenergetic beams of alkali metal atoms and iodomethane molecules; D. R. Hershbach, *Discussions Faraday Soc.*, **33**, 149 (1962).

(10) H. W. Melville, J. C. Robb, and R. C. Tutton, *ibid.*, **14**, 150 (1953).

Results

The rate of disappearance of the stable free radical, 1-ethyl-4-carbomethoxy-pyridinyl (1), in the presence of halocarbons can be followed very conveniently by means of a spectrophotometer, either in the visible region (near 6300 Å.) or in the ultraviolet (at 3950 or 3100 Å.).^{2a} The procedure varied somewhat with the halocarbon used (see Experimental) but basically involved the introduction of separate solutions of free radical and halocarbon into an apparatus, temperature equilibration, mixing, and measurement. The most important experimental errors arose from the measurement of the concentration of the halocarbon and the control of the temperature, limiting the accuracy of the rate constants in most cases to ±5-7%. The rate constants listed below are corrected (factor of 2 for the pseudo-unimolecular reactions) for the fact that two pyridinyl radicals disappear for each halogen atom abstracted.¹³

The rate of reaction of 1 with dibromomethane (2) was examined in a number of solvents, ranging in polarity from dichloromethane (3, Z 64.2)¹⁴ through acetonitrile (Z 71.3) and isopropyl alcohol (Z 76.3)¹⁵ to ethanol (Z 79.6).¹⁵ The rate constants were insensitive to the solvent polarity, ranging from 0.3 × 10⁻⁴ l. mole⁻¹ sec.⁻¹ in isopropyl alcohol to 1.0 × 10⁻⁴ l. mole⁻¹ sec.⁻¹ in acetonitrile. The rate constants listed in Table I are for the elementary reaction shown in eq. 4.



To ascertain the effect of bond strength on the rate of the atom transfer reaction, the rates of reaction of 1 with the series dichloromethane (3), bromochloromethane (4), and iodochloromethane (5) were examined. The great sensitivity of the reaction to the nature of the halogen demonstrates that the reaction

(11) M. Szwarc in "The Transition State," Special Publication No. 16, The Chemical Society, London, 1962, p. 101.

(12) C. Walling, "Free Radicals in Solution," John Wiley and Sons, Inc., New York, N. Y., 1957.

(13) These constants therefore differ by a factor of 2 from those reported in the preliminary communication in which only the measured second-order constants were given; E. M. Kosower and I. Schwager, *J. Am. Chem. Soc.*, **86**, 4493 (1964).

(14) E. M. Kosower, *ibid.*, **80**, 3253 (1958).

(15) The Z-value for the halocarbon-alcohol solutions used is somewhat lower than that for the pure solvents; see the notes to Table I.

TABLE I
RATES OF REACTION OF 1-ETHYL-4-CARBOMETHOXYPYRIDINYL
WITH DIBROMOMETHANE^a

Radical concn. × 10 ⁻³ , <i>M</i>	Halocarbon concn., <i>M</i>	Solvent	Temp., ^b °C.	<i>k</i> _{obsd} × 10 ⁴ , sec. ⁻¹	<i>k</i> ₂ ^{c,d} × 10 ⁴ , l. mole ⁻¹ sec. ⁻¹
1.24	1.30	CH ₃ CN	25	5.33	1.03
0.25	0.98	CH ₃ CN	25 ^h	3.61	0.93
1.33	.84	CH ₃ CN	25	3.43	1.02
1.11	.41	CH ₃ CN	25	1.52	0.93
0.98	.84	CH ₂ Cl ₂ ^e	25	1.60	.48
1.93	.84	<i>i</i> -PrOH	25	0.94 ^f	.28
0.24	1.28	EtOH	25	ca. 1.7 ^g	ca. .33
.51	0.98	CH ₃ CN	30 ^h	5.77	1.48
.60	0.97	CH ₃ CN	35 ^h	9.90	2.55

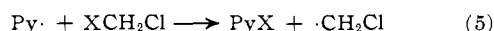
^a Radical solutions in acetonitrile decrease in concentration at 75° with an estimated half-life of 12 days. ^b No more than ±1°. The quartz cell (apparatus A) in which the reactions were run was thermostated to ±0.2°, but the remainder of the apparatus was exposed (in the closed cell compartment) to a temperature as much as 1–2° higher. ^c The rate constants obtained from *k*_{obsd} by dividing by the halocarbon concentration are corrected by dividing by the number of reactive halogens (*two* for dibromomethane) and by the number of pyridinyl radicals consumed for every molecule of reacting halocarbon (*two*). ^d Estimated reliability 5–7%. ^e *Z* 64.2. ^f Corrected for the solvent reaction with radical, *k*₁ = 2.61 × 10⁻⁵ sec.⁻¹ for 0.84 *M* CH₂Cl₂ in isopropyl alcohol (*Z* 75.5). A solution of dibromomethane (0.84 *M*) in isopropyl alcohol has *Z* 75.8. ^g Estimated after an approximate correction for the reaction of solvent with radical. ^h Carried out in a sealed cell; temperature ±0.2°.

TABLE II
RATES OF REACTION OF 1-ETHYL-4-CARBOMETHOXYPYRIDINYL
WITH HALOCHLOROMETHANES IN ACETONITRILE

Radical concn. × 10 ⁻² , <i>M</i>	Halocarbon concn., <i>M</i>	Temp., ^a °C.	<i>k</i> _{obsd} , sec. ⁻¹	<i>k</i> ₂ ^b , l. mole ⁻¹ sec. ⁻¹
1.48	CICH ₂ Cl (0.90)	80	18.2 × 10 ^{-6c}	5.0 × 10 ⁻⁶
0.92	(.85)	75	8.7 × 10 ^{-6c}	2.6 × 10 ⁻⁶
1.64	(.78)	70	9.6 × 10 ^{-6c}	3.1 × 10 ⁻⁶
1.38	(.78)	65	3.34 × 10 ^{-6c}	1.1 × 10 ⁻⁶
1.08	(.90)	60	1.79 × 10 ^{-6c}	0.50 × 10 ⁻⁶
0.65	(.85)	50	1.06 × 10 ^{-6c}	0.31 × 10 ⁻⁶
...	...	25	...	2.6 ± 1.0 × 10 ⁻⁶
1.25	BrCH ₂ Cl (0.84)	25	8.57 × 10 ⁻⁶	5.1 × 10 ⁻⁶
1.51	(1.18)	25	1.12 × 10 ⁻⁴	4.8 × 10 ⁻⁵
1.72	(1.15)	30	2.38 × 10 ⁻⁴	1.04 × 10 ⁻⁴
0.65	(1.18)	35	3.68 × 10 ⁻⁴	1.56 × 10 ⁻⁴
1.90	ICH ₂ Cl (0.0095)	25 ^{a,f}	...	1.6 × 10 ⁻¹
1.72	(0.0086)	25 ^{a,f}	...	1.1 × 10 ⁻¹
1.20	ICH ₃ (2.96)	25	2.95 × 10 ⁻⁵	5.0 × 10 ⁻⁶
0.96	(2.89)	25	2.56 × 10 ⁻⁵	4.4 × 10 ⁻⁶
1.87	(3.12)	30	4.13 × 10 ⁻⁵	6.6 × 10 ⁻⁶
1.96	(2.89)	35	6.67 × 10 ⁻⁵	1.15 × 10 ⁻⁵
0.82	(3.12)	20	1.33 × 10 ⁻⁵	2.1 × 10 ⁻⁶

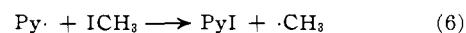
^a ±0.2° except as noted. ^b See footnotes *c* and *d*, Table I. ^c Rates given for dichloromethane are *initial rates*. The apparent rates slowly decrease owing to the formation of a product (or products) which have some absorption in the region at which the loss of radical is followed. ^d Extrapolated from the data at 80°, *E*_a = 22 or 26 kcal./mole, the range suggested by the plot of log *k* vs. 1/*T*. ^e See footnote *b*, Table I. ^f Determined under bimolecular conditions. Data treated according to the case for consecutive second-order reactions, A + B → C + D(*k*₁), A + C → F + G(*k*₂), in which *k*₂ ≫ *k*₁; cf. ref. 16, p. 181.

follows eq. 5 and that the leaving group in each case is the chloromethyl radical. In the case of dichloromethane, the reaction was far too slow to be measured at 25°; in fact, it was used as a low-polarity solvent



for the reaction of 1 with 2. However, the stability of acetonitrile solutions of 1 at elevated temperatures was sufficient to permit the reaction with dichloro-

methane to be followed. The rate data are shown in Table II. Rate data for iodomethane (6) are also included (eq. 6).



To examine the effect of chlorine substitution on the rate of chlorine atom transfer, the rates of reaction of 1 with trichloromethane (7) and tetrachloromethane (8) were measured. The results are given in Table III along with the extrapolated value for 3, comprising

TABLE III
RATES OF REACTION OF 1-ETHYL-4-CARBOMETHOXYPYRIDINYL
WITH POLYCHLOROMETHANES IN ACETONITRILE

Radical concn. × 10 ⁻² , <i>M</i>	Halocarbon concn., <i>M</i>	Temp., ^a °C.	<i>k</i> _{obsd} × 10 ⁴ , sec. ⁻¹	<i>k</i> ₂ ^b , l. mole ⁻¹ sec. ⁻¹
...	CH ₂ Cl ₂	25	...	2.6 × 10 ⁻⁸
2.14	CHCl ₃ (1.43)	25	1.75	2.04 × 10 ⁻⁵
0.48	(1.33)	25	1.70	2.14 × 10 ⁻⁵
1.65	(1.33)	35	4.62	5.8 × 10 ⁻⁵
0.93	(1.33)	15	0.583	7.3 × 10 ⁻⁶
2.41	CDCl ₃ (1.49)	25	1.56	1.75 × 10 ^{-5d}
1.91	(1.55)	25	1.57	1.68 × 10 ^{-5d}
0.013	CCl ₄ (0.00065)	25 ^f	...	3.9 ^g

^a ±0.2° except as noted. ^b See footnotes *c* and *d*, Table I. ^c From Table II. ^d Average *k*(CHCl₃)/average *k*(CDCl₃) = *k*_H/*k*_D = 1.21. ^e It is unlikely that pyridinyl radical abstraction of chlorine from the intermediate trichloromethyl radical could compete with the (probably) diffusion-controlled recombination of the pyridinyl radical and the trichloromethyl radical. It may be estimated from the second dissociation energy of tetrachloromethane (H. A. Skinner and G. Pilcher (*Quart. Rev.* (London), 17, 268 (1963)) that the abstraction reaction would have a total rate constant (for 3 chlorines) of less than 10⁴ l. mole⁻¹ sec.⁻¹. ^f See footnote *f*, Table II. ^g Corrected by dividing the measured *k*₂ by the statistical factor of 4.

the series: CICH₂Cl, ClCHCl₂, and ClCCl₃. That chlorine atom transfer is the reaction under study is shown by the magnitude of the secondary isotope effect, *k*_H/*k*_D, for the reaction of 1 with trichloromethane and deuteriotrichloromethane (7D). Data for 7D are also listed in Table III.

The rate data (optical density (*D*) as a function of time) were treated either by plotting log *D* against time or by the method of Guggenheim.¹⁶ The latter method was especially useful when small residual absorptions attributed to by-products made an accurate infinity determination difficult. In most runs, however, absorption of the radical 1 decreased to zero. Typical plots are presented in Fig. 1 (iodomethane), 2 (dibromomethane), and 3 (tetrachloromethane). Tetrachloromethane and iodochloromethane were much too reactive to study under the usual conditions involving a large excess of the halocarbon; their reactivities were followed under bimolecular conditions, yielding for 8 the excellent plot shown in Fig. 3.

The activation energies for the reaction of 1 with five halocarbons were measured (Table IV) but did not reveal any well-defined pattern.

Products.—In six instances—for the halocarbons, 2, 4, 5, and 8—the solvent was removed from the reaction solution at the termination of the rate measurement, *n*-hexane was introduced (except for 8 for which the acetonitrile distillate was examined directly), the soluble material removed, and its spectrum measured. Absorption was largely between 2100 and 2900

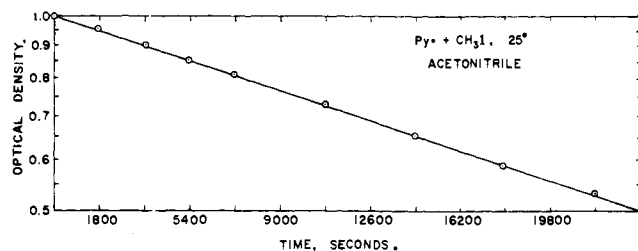


Fig. 1.—A plot of optical density *vs.* time for the reaction of 1-ethyl-4-carbomethoxy-pyridinyl with iodomethane in acetonitrile at 25° as followed at 6325 Å.

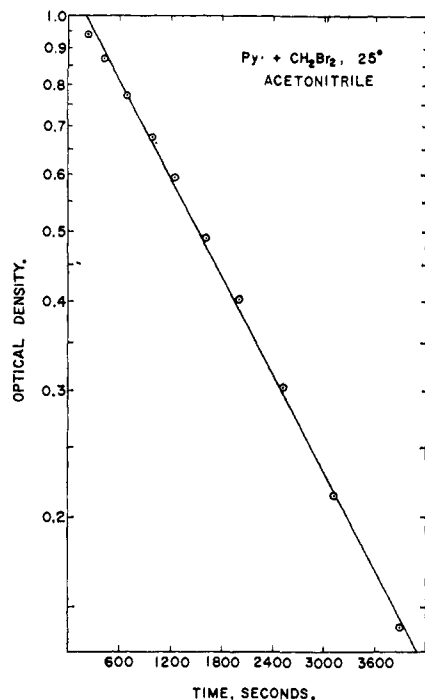


Fig. 2.—A plot of optical density at 6325 Å. *vs.* time for the reaction of 1-ethyl-4-carbomethoxy-pyridinyl with dibromomethane in acetonitrile at 25°.

Å. with very much weaker broad maxima sometimes found between 3600 and 3700 Å. For **2**, λ_{\max} was 2420 Å. (minimum ϵ 10,000, for theoretical yield of 50%) and for **8**, λ_{\max} was 2380 Å. (minimum ϵ 20,000). The spectra also had either shoulders or rising absorp-

TABLE IV
ACTIVATION ENERGIES FOR THE REACTION OF 1-ETHYL-4-CARBOMETHOXY-PYRIDINYL WITH HALOCARBONS IN ACETONITRILE

Halocarbon	Temp. range, °C.	E_a , kcal./mole
CH ₂ Cl ₂	50-80	24 ± 2
CH ₂ BrCl	25-35	22 ± 2
CHCl ₃	15-35	19 ± 1
CH ₃ I	20-35	19 ± 2
CH ₂ Br ₂	25-35	18 ± 2

tion in the region from 2800 to 2900 Å. These facts are consistent with the presence of either 1,2- (**9**) or 1,4-dihydropyridines (**10**). Simple derivatives of the former structure have not been characterized, but the spectrum of 1,4,4-trimethyl-1,4-dihydropyridine and its other properties are well known.¹⁷ Its maxima are at 2708 Å. (ϵ 3200) and 2305 Å. (ϵ 7500) in isooctane. The volatility of the hexane-soluble material (which

(17) E. M. Kosower and T. S. Sorensen, *J. Org. Chem.*, **27**, 3764 (1962).

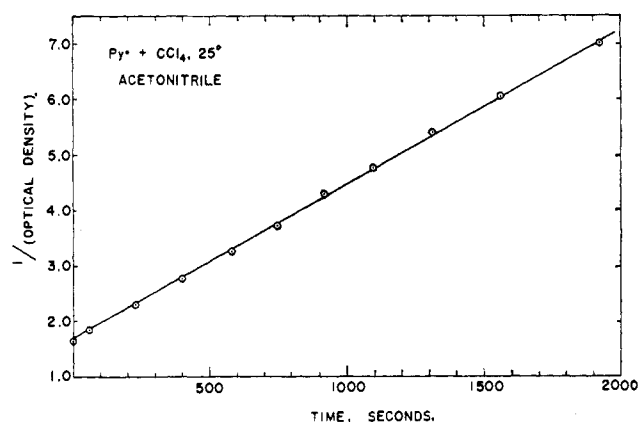
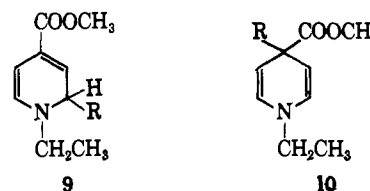


Fig. 3.—A plot of the reciprocal of optical density at 3950 Å. *vs.* time for the reaction of 1-ethyl-4-carbomethoxy-pyridinyl with tetrachloromethane in acetonitrile at 25°.

codistilled with the solvent) and its air sensitivity (the spectrum changed upon exposure to oxygen) are in agreement with the presence of **9** or **10**.

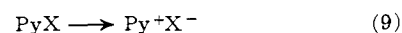
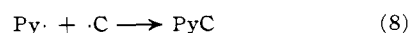
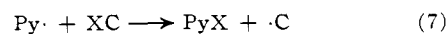


After removal of the hexane-soluble material, the residue was dissolved in degassed water and its ultra-violet spectrum measured. It was virtually identical with that of the 1-ethyl-4-carbomethoxy-pyridinium ion, with maxima at 2740 and 2200 Å.¹⁸ Given the known absorption coefficient (ϵ 4550 at 2735 Å.), the yield of pyridinium ion was 51 ± 4%. The spectrum did not change upon exposure to air and the solution contained halide ion (precipitate with silver ion).

In summary, the products of the reaction of 1-ethyl-4-carbomethoxy-pyridinyl with halocarbons are 50% 1-ethyl-4-carbomethoxy-pyridinium ion and *ca.* 50% dihydropyridines. It will be of some interest to examine the nature of the latter for different halocarbon substrates as a part of a study on how radical **1** reacts with other radicals.

Discussion

The reaction of the pyridinyl radical **1** with halocarbons results in the abstraction of a halogen atom and the formation of a covalently bonded adduct of the halide ion and the pyridinium ion along with a radical derived from the halocarbon. The adduct rapidly ionizes to give the pyridinium halide while the halocarbon radical reacts with another molecule of the pyridinyl radical, presumably in a diffusion-controlled reaction, to form a dihydropyridine (eq. 7-9).



The evidence in support of this mechanism is: (a) the formation of pyridinium halide in the expected yield (*ca.* 50%); (b) the extreme sensitivity of the reaction rate to the nature of the halogen atom being

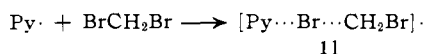
(18) E. M. Kosower, "Molecular Biochemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1962, p. 185.

abstracted; (c) the small isotope effect, k_H/k_D (1.21), observed by a comparison of **7** and **7D**; (d) the dihydropyridines which are formed; (e) the extremely small solvent effect on the rate, demonstrating that the transition state cannot differ greatly from the initial state in charge separation.

The isotope effect found for the reaction of trichloromethane (**7**) and **7D** with **1** appears to be somewhat greater than that found by Zavitsas and Seltzer¹⁹ for the formation of methyl and trideuteriomethyl radical from methyltrideuteriomethylphenylcarbonyl hypochlorite for which $k_H/k_D = 1.14 \pm 0.01$ at 20°.

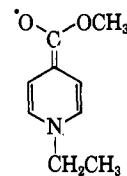
Solvent Effect.—The effect of solvent polarity on the rate of a reaction is one of the most useful criteria of mechanism. If any appreciable difference in charge separation exists between the initial state and the transition state, the rate of the reaction will change in a manner dependent upon the magnitude of the solvent polarity change. A convenient measure of solvent polarity, Z -value,¹⁴ can be used to evaluate the extent of the change in charge separation. The rate of solvolysis of *t*-butyl chloride changes by a factor of over 3×10^5 for the solvent change from ethanol (Z 79.6) to water (Z 94.6). Although part of the rate change is caused by a rise in the ground-state free energy of the *t*-butyl chloride (perhaps 10%) in water, the major effect can be ascribed to the increased charge separation in the transition state for ionization. Clarke and Taft²⁰ have estimated the change in the charge on carbon (and chlorine) by means of salt effects as $q = 0.2 \rightarrow q = 0.8$. Another estimate made by a method similar to that used to estimate the excited-state dipole moments of α,β -unsaturated ketones²¹ by the use of Z -values gives a result in agreement with that derived from salt effects.²² These figures provide us with some perspective for the consideration of changes in charge separation in radical reactions.

The rate constant for bromine abstraction by radical **1** from dibromomethane shows little variation over a solvent polarity range of $\Delta Z = 15.4$, from dichloromethane (Z 64.2) to ethanol (Z 79.6). In fact, the rate constant is slightly smaller in isopropyl alcohol and ethanol than in dichloromethane (*ca.* 0.3×10^{-4} l. mole⁻¹ sec.⁻¹ vs. 0.5×10^{-4} l. mole⁻¹ sec.⁻¹) and rises to 1.0×10^{-4} l. mole⁻¹ sec.⁻¹ in acetonitrile. Compared to the factor of 10^5 which one might have expected for this solvent polarity range for a reaction in which the change in charge separation in the transition state as compared to the initial state is comparable to *t*-butyl chloride, these rate effects must be regarded as negligible. The transition state for bromine abstraction must thus possess the same degree of charge separation as the initial state, and is best expressed as **11**.



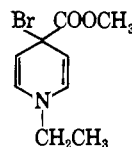
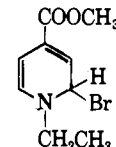
The solubility of the radical **1** in *n*-heptane is considerable, *ca.* 0.02 *M*. The molecule cannot have a very large dipole moment. On the basis of the infrared spectrum²³ which exhibits a carbonyl frequency greatly reduced in comparison to that of 1-ethyl-4-

carbomethoxypyridinium ion or methyl isonicotinate, form **12** can be written as an important contribution to the structure of the ground state of **1**. The dipole

**12**

moment expected for a molecule like **1** with a large contribution from **12** is probably somewhat greater than that of an ether (1.2 D.). Dibromomethane also has a modest dipole moment (1.5 D.).²³ The net dipole moment of the initial state could be between 0 and 3 D. The transition state **11** must also have a dipole moment between 0 and 3 D. in view of the negligible solvent effect upon the rate. "Carbonium ion stabilization" of the transition state **11** should have been unusually effective because the form which expresses this contribution includes a pyridinium ion, *i.e.*, a stable "carbonium ion." The solvent effect conclusively demonstrates that such stabilization cannot be important in the case of pyridinyl radicals whatever role it might play in other cases.²⁴⁻³⁰ This matter will be discussed further below.

The transition state **11** must proceed to products with a low net dipole moment, or, more precisely, with total charge separation not very different from that of **11**. The only reasonable possibility is that the bromomethyl radical is formed along with the covalently bonded adduct of bromide ion with the pyridinium ring, either **13** or **14**. The bromide adduct of 1-(2,6-dichlorobenzyl)-3,5-dicarbamidopyridinium ion has

**13****14**

been reported³¹ and 1-methyl-3,5-dicyanopyridinium iodide appears to exist entirely in the form of a covalent adduct in most solvents.³² One can presume that adduct is simply less stable in the case of the 1-ethyl-4-carbomethoxypyridinium ion and would readily ionize after formation to yield the ionic pyridinium halide.

The opinion that "carbonium ion stabilization" is responsible for the effects of polar groups (*e.g.*, *p*-methoxy, *p*-nitro, etc.) on the rates of radical reactions is based almost entirely upon the correlation of the rates with the Hammett σ^+ -constant.²⁴⁻³⁰

(23) Cf. A. A. Maryott, M. E. Hobbs, and P. M. Gross, *J. Am. Chem. Soc.*, **63**, 659 (1941), who cite 1.43 D. for the gas-phase value.

(24) E. S. Huyser, *ibid.*, **82**, 394 (1960).

(25) E. S. Huyser, H. Schimke, and R. L. Burham, *J. Org. Chem.*, **28**, 2141 (1963).

(26) Reference 12, p. 158.

(27) C. H. Bamford, A. D. Jenkins, and R. Johnston, *Trans. Faraday Soc.*, **55**, 418 (1959).

(28) G. A. Russell and R. C. Williamson, Jr., *J. Am. Chem. Soc.*, **86**, 2357 (1964).

(29) D. G. Hendry and G. A. Russell, *ibid.*, **86**, 2368 (1964).

(30) P. D. Bartlett and C. Ruchardt, *ibid.*, **82**, 1760 (1960).

(31) K. Wallenfels and H. Schüly, *Ann.*, **621**, 111 (1959).

(32) M. Fischer, unpublished results.

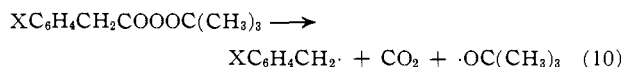
(19) A. A. Zavitsas and S. Seltzer, *J. Am. Chem. Soc.*, **86**, 1265 (1964).

(20) G. A. Clarke and R. W. Taft, *ibid.*, **84**, 2295 (1962).

(21) E. M. Kosower and D. C. Remy, *Tetrahedron*, **5**, 281 (1959).

(22) E. M. Kosower, unpublished results. However, cf. R. E. Robertson, *Can. J. Chem.*, **42**, 1707 (1964).

The most striking aspect of these correlations³³ is the fact that the ρ -values are all small (from 0 to -1.46 except for the -4.54 reported for the reaction of diphenylpicrylhydrazyl with phenols³⁴) and *negative*. To take one of the most interesting and best studied examples, the concerted decomposition of *t*-butyl phenyl peracetates (eq. 10),^{30,35,36} Bartlett and Rü-



chardt have reported a ρ of -1.2 .

The ionization potentials of benzyl radicals³⁷ are extremely sensitive to the nature of the substituent ($\rho = -19.2$), and it is clear that the *p*-methoxybenzyl radical (I_p 6.82 v.) is much less stable than the benzyl radical (I_p 7.76 v.) or the *p*-cyanobenzyl radical (I_p 8.36 v.). We should thus expect that a transition state leading to a benzyl radical (eq. 10) would be stabilized by electron-attracting groups and destabilized by electron-supplying groups. This supposition finds confirmation in the ready preparation of radical anions of nitroaromatics and other aromatic compounds bearing electron-attracting substituents. The experimental results cited above show clearly that the substituents have effects opposite to what our expectations are. Semenov³⁸ has offered a simple solution to the apparent contradiction by pointing out that the substituents could readily affect ground-state bond energies to the extent necessary to account for the small ρ -values observed. In these terms, electron-attracting groups strengthen the bond which is to be broken, and electron-supplying groups weaken that bond. The reason for these small variations in bond strength (which might be larger than the observed rate effects if we allow for the stabilization or destabilization (correlatable by ρ) expected in the transition state) might be as simple as the electrostatic repulsion (or attraction) model proposed originally for substituent effects on radical reactions by Swain, Stockmayer, and Clarke,³⁹ or might be the combination of a number of more subtle effects.

If a variation in ground-state bond energies were responsible for a substantial part of the polar substituent effect on radical reactions, we would not expect a solvent polarity effect on the reaction rates since there would then be no reason to suppose that the charge separation in the transition state would be any greater than that in the initial state. Indeed, Bartlett and Lee⁴⁰ have found that the rate of reaction of *t*-butyl *p*-methoxyphenylperacetate is 3.8 times higher in ethanol than in cyclohexane (and 3.6 times higher in nitrobenzene) at 70.4° in the presence of $0.1 M$ styrene to inhibit chain decomposition. The ΔZ for cyclohexane and ethanol is 19.5, and only a reaction with a transition state very similar in charge separation to the initial state could be so insensitive to such a large change in solvent polarity.

(33) Cf. also W. K. Wilmarth and N. Schwartz, *J. Am. Chem. Soc.*, **77**, 4543, 4551 (1955).

(34) J. C. McGowan, T. Powell, and R. Raw, *J. Chem. Soc.*, 3103 (1959).

(35) P. D. Bartlett and R. R. Hiatt, *J. Am. Chem. Soc.*, **80**, 1398 (1958).

(36) P. D. Bartlett and D. M. Simons, *ibid.*, **82**, 1753 (1960).

(37) A. G. Harrison, P. Kebarle, and F. P. Lossing, *ibid.*, **83**, 777 (1961).

(38) N. N. Semenov, "Some Problems in Chemical Kinetics and Reactivity," Princeton University Press, Princeton, N. J., 1958, p. 83.

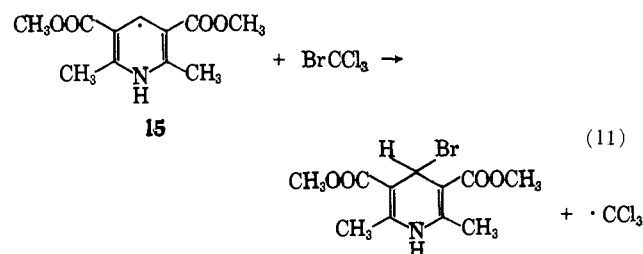
(39) C. G. Swain, W. H. Stockmayer, and J. T. Clarke, *J. Am. Chem. Soc.*, **72**, 5426 (1950).

(40) Cited in ref. 30, footnote 16.

Cook and Norcross⁴¹ found that a factor of less than 2 separated the rates of disproportionation of 2,6-di-*t*-butyl-4-isopropylphenoxy, a moderately stable free radical, in chlorobenzene and benzonitrile. The ΔZ for these two solvents would probably be less than 10. Although these authors found an "isokinetic relationship" for the solvents they used in the manner suggested by Leffler,⁴² the experimental errors in the energies and entropies of activation are probably greater than indicated, and Ross and his co-workers⁴³ have severely criticized this approach to solvent effects on the grounds that the relationships only reflect experimental errors.

Recently, Hendry and Russell²⁹ have attempted to measure the effect of solvent polarity on the hydrogen abstraction reactions of peroxy radicals. Their rate data are a composite of a number of elementary steps, and the meaning of the results is obscure. A macroscopic measure of solvent polarity, the dielectric constant, was used to evaluate the data. A factor of slightly more than 3 separated the rates of the fastest and slowest reactions. The solvent effect is small enough ($\Delta Z = 11$) and the scatter of the data great enough (as plotted against a function of dielectric constant) so that one may conclude that the transition state does not have any greater charge separation than the initial state.

Kurz, Hutton, and Westheimer⁴⁴ discovered that certain 1,4-dihydropyridines reacted either thermally or photochemically with tetrachloromethane (**8**) or bromotrichloromethane in a chain reaction which must involve pyridinyl radicals. We have previously shown that the reaction of a pyridinyl radical (**1**) with **8** does occur,^{2a} substantiating the propagation reaction proposed for the dihydropyridines. We can now further assert, on the basis of the transition state description we have developed for the reaction of **1** with **2**, that the propagation step leads to an intermediate in which halide ion is covalently bonded to the pyridinium ring, the second of the two alternatives (eq. 11) suggested by Kurz, *et al.*⁴⁴ Furthermore, we can estimate the rate of bromine abstraction reaction (eq. 11) as *ca.* 10^5 l. mole⁻¹ sec.⁻¹ from the data in Tables II and III and the assumption that the radical **15** will be ten times as reactive as **1**.⁴⁵



Nothing is known about the initiation reaction which leads to radical **15**.⁴⁴ A possible source is the complex between the dihydropyridine and bromotrichloromethane which could undergo electron transfer on

(41) C. D. Cook and B. E. Norcross, *J. Am. Chem. Soc.*, **81**, 1176 (1959).

(42) M. C. Alder and J. E. Leffler, *ibid.*, **76**, 1425 (1954).

(43) R. C. Petersen, J. H. Markgraf, and S. D. Ross, *ibid.*, **83**, 3819 (1961).

(44) J. L. Kurz, R. Hutton, and F. H. Westheimer, *ibid.*, **83**, 584 (1961).

(45) The relative rate of halogen abstraction per halogen atom from bromotrichloromethane and tetrachloromethane by methyl radical favors the former reaction by a factor of 6700; Table 5 in M. Szwarc, ref. 11.

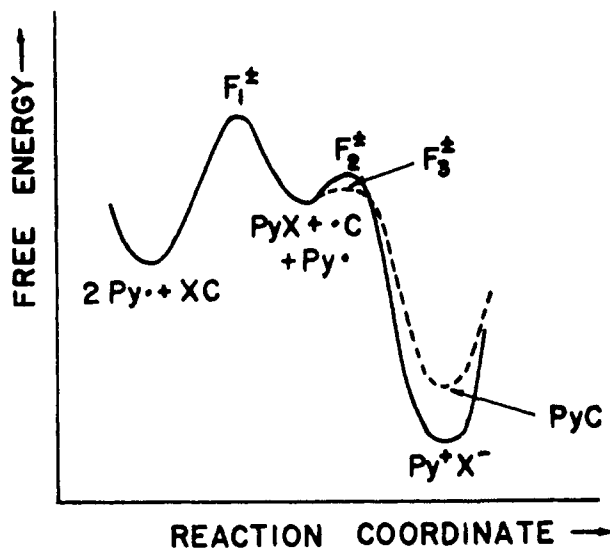


Fig. 4.—A free-energy *vs.* reaction-coordinate diagram for the reaction of pyridinyl radicals with halocarbons. $\text{Py}\cdot$ is the pyridinyl radical (in this case, 1-ethyl-4-carbomethoxypyridinyl), XC is the halocarbon, PyX is the covalently bonded adduct of halide ion to the pyridinium ring, $\text{C}\cdot$ is the radical produced by abstraction of the halogen X from the halocarbon, Py^+X^- is the ionic pyridinium halide, and PyC is the dihydropyridine (or dihydropyridines) formed by combination of the pyridinyl radical with the halocarbon radical.

light absorption, followed by dissociation to radicals. Aliphatic amines, which are presumably weaker donors than the dihydropyridines, react photochemically with bromotrchloromethane and other halocarbons to generate trichloromethyl radical.⁴⁶ However, the intervention of traces of ions such as cuprous ion cannot be excluded.⁴⁷

Dittmer and Fouty⁴⁸ have described a number of halogen abstraction reactions which probably proceed through pyridinyl radicals. Typically, 1-benzyl-3-carbamidopyridinyl reacted with hexachloroacetone to produce the pentachloroacetyl radical. The source for the pyridinyl radical was 1-benzyl-3-carbamido-1,4-dihydropyridine and again the initiation process was obscure.

Interpretation of Rate Effects.—A free energy *vs.* reaction coordinate diagram is shown in Fig. 4 in order to illustrate the relationships implied in the reaction mechanism suggested for 1 and the halocarbons.

The free energies of dissociation of the carbon-chlorine bond in the compounds dichloromethane (3), trichloromethane (7), and tetrachloromethane (8) can be derived from the rate data with the aid of a simple assumption. Consider the free energy *vs.* reaction coordinate diagram in Fig. 5. The difference in the rate of reaction of 1 with 7 and 8 corresponds to the sum of ΔF_1 and ΔF_2 . If we assume that $\Delta F_2 = \Delta F_3$, and further that the intervention of PyCl does not change the relationship between the two processes, of eq. 12a and 12b, then the rate difference is the same as might have been expected from the difference in dissociation energies were they available. The two factors which control ΔF_{12} and ΔF_{13} are the bond energies and the stabilities of the product radicals.

(46) D. P. Stevenson and G. M. Coppinger, *J. Am. Chem. Soc.*, **84**, 149 (1962).

(47) M. Asscher and D. Vofsi, *J. Chem. Soc.*, 2261 (1961).

(48) D. C. Dittmer and R. A. Fouty, *J. Am. Chem. Soc.*, **86**, 91 (1964).

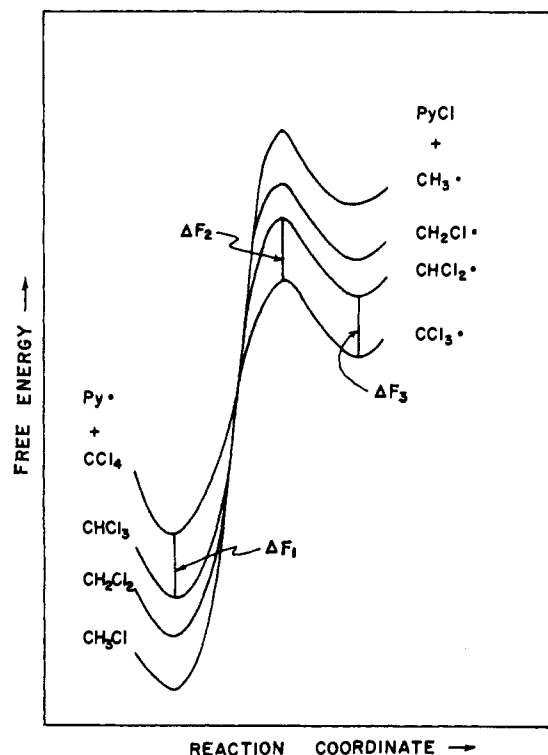
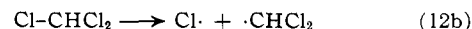
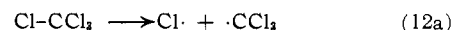
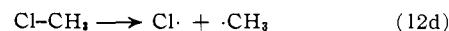
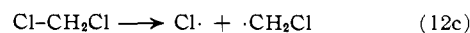


Fig. 5.—A free-energy *vs.* reaction-coordinate diagram for the reaction of pyridinyl radical (in this case, 1-ethyl-4-carbomethoxypyridinyl) with the chloromethane series: CH_3Cl , CH_2Cl_2 , CHCl_3 , and CCl_4 . The product PyCl represents the covalently bonded adduct of chloride ion to the pyridinium ring. The curves are drawn so as to divide the measured rate differences about equally between bond energy differences and product radical stability. It must be emphasized that this division is arbitrary and that all of the differences may well reside in, for example, product radical stability.



An estimate of the rate constant may be made for chloromethane (eq. 12d) by a comparison of the data for dichloromethane and iodochloromethane with that for iodomethane. The rate data for the chloro-



methanes are summarized in Table V. These suggest that the difference between the dissociation energies for the carbon-chlorine bond in chloromethane and

TABLE V
RATE AND FREE-ENERGY COMPARISON OF CHLOROMETHANES

Compound	$k_3(25^\circ)^a$	ΔF_{12}^* , kcal./mole	$\Delta\Delta F_{12}^*$, kcal./mole
CH_3Cl	1.1×10^{-12}	31.4	
CH_2Cl_2	2.6×10^{-8}	25.5	5.9
CHCl_3	2.1×10^{-5}	21.6	3.9
CCl_4	3.3	14.5	7.1

^a Per chlorine atom in l. mole⁻¹ sec.⁻¹.

that in tetrachloromethane is 16.9 kcal./mole. The dissociation energies reported by Cottrell⁴⁹ are $D(\text{CH}_3-\text{Cl}) = 80$ kcal./mole and $D(\text{CCl}_3-\text{Cl}) = 68 \pm 3$ kcal./mole, yielding 9–15 kcal./mole for the difference, in poor agreement with the difference derived from our kinetic data. Cottrell⁴⁹ cites $D(\text{CHCl}_2-\text{Cl})$

(49) T. L. Cottrell, "The Strengths of Chemical Bonds," 2nd Ed., Butterworths, London, 1958, p. 203.

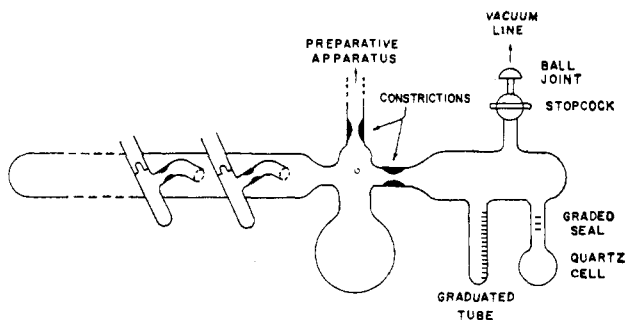


Fig. 6.—An apparatus for the preparation of multiple samples of a solution of 1-ethyl-4-carbomethoxy-pyridinyl of known concentration. The actual apparatus had twelve breakoff seal storage tubes, six to each side.

= 72 kcal./mole or less as doubtful, but our kinetic data suggest 70.2 kcal./mole, assuming the value of 80 kcal./mole for chloromethane. We may also derive the value of 74.1 kcal./mole for $D(\text{CH}_2\text{Cl}-\text{Cl})$. Papulov⁵⁰ lists $D(\text{CH}_3-\text{Cl}) = 83.5$ kcal./mole and $D(\text{CCl}_3-\text{Cl}) = 68.4$ kcal./mole, for a difference of 15.1 kcal./mole.

It is, of course, difficult to decide whether or not the rate effects (that is, the effect of chlorine substitution on the rate of the halogen abstraction by the radical 1) are due to bond energy differences or to differences in the stabilities of the product radicals. Assuming that the latter effect is completely responsible for the rate effects, one may estimate stabilization energies for the product radicals relative to methyl radical as: $\cdot\text{CH}_2\text{Cl}$, 6 kcal./mole; $\cdot\text{CHCl}_2$, 10 kcal./mole; and $\cdot\text{CCl}_3$, 17 kcal./mole. On the other hand, it might be anticipated that the carbon-chlorine bond in tetrachloromethane would be weaker than the same bond in chloromethane owing to van der Waals repulsion.⁵¹

The series of compounds, XCH_2Cl , is interesting because the "leaving group" is constant. However, the rate differences must be ascribed to the differences between the C-X bond energies in the halochloromethanes and the covalent adduct of the halide ion to the pyridinium ring, PyX . If one were to prepare a free-energy *vs.* reaction-coordinate diagram like that of Fig. 5, and take $\Delta F_2 = \Delta F_3$, the rate difference would be equivalent to $\Delta F_1 - \Delta F_2$.

The more reactive the radical engaged in the act of abstraction, the less likely it is to discriminate among halocarbons. This statement is equivalent to the proposition that less bond breaking is required in the transition state with a reactive radical than with an unreactive radical. We could then expect that 1-phenylethyl radical and sodium atom would be less selective with regard to trichloromethane and tetrachloromethane than the pyridinyl radical. According to the data summarized by Walling,⁵² the transfer constants ($k_{\text{transfer}}/k_{\text{polymerization}}$) for a 1-phenylethyl radical (*i.e.*, a polystyryl radical) for tetrachloromethane and trichloromethane are in the ratio of 180:1, while the corresponding ratio for sodium atom is only 4:1. In contrast, pyridinyl radical 1 reacts with 8 about 10^5 faster than with 7. On an energy scale, the discrimination ratio, pyridinyl:1-phenylethyl:so-

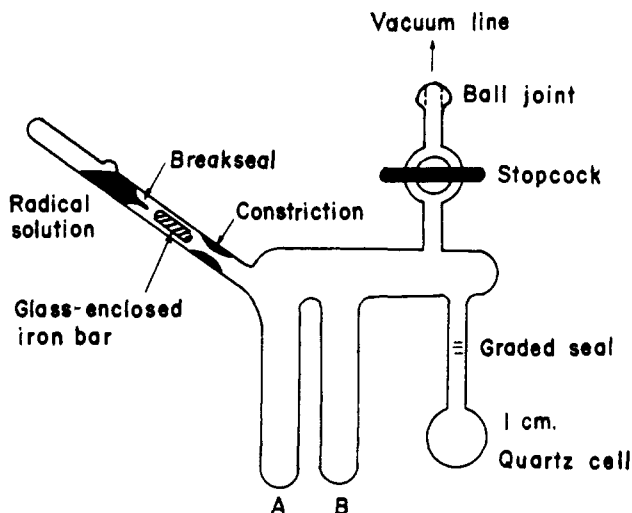


Fig. 7.—An apparatus for the determination of the rates of reaction of 1-ethyl-4-carbomethoxy-pyridinyl with halocarbons. The precise procedure may be varied to fit the experiment, but, typically, the breakseal is broken and the radical solution is introduced on the left. The side arm and breakoff seal tube are sealed off. The radical solution is kept frozen with Dry Ice-acetone (195°K.) and a halocarbon solution of the appropriate concentration is distilled into the other tube with liquid nitrogen (77°K.) as coolant. Both tubes are brought to the desired temperature in a bath, the solutions rapidly mixed, the reaction mixture poured into the quartz cell, and the apparatus introduced into the Cary Model 14 spectrophotometer. Measurements were usually begun from 1 to 2 min. after mixing. If the reactions were slow, the quartz cell could be replaced by a Pyrex cell made from square tubing, the latter cooled in liquid nitrogen to freeze the contents, and sealed off.

dium, is 9.9:3.8:1. Valuable information about radical reactivity should become available by measurement of discrimination ratios against such compound pairs as 8 and 7.

Experimental

Materials. 1-Ethyl-4-carbomethoxy-pyridinium Iodide, M.p. 111–112°.^{2a}—All halocarbons were fractionated before use and their purity checked with g.l.c. (gas-liquid phase chromatography). In a number of cases (*e.g.*, trichloromethane), the halocarbon was poured through a neutral alumina column before fractionation. A spinning band column (platinum band, Nester and Faust) was used. All solvents except acetonitrile were fractionated before use and almost all were Spectrograde. Bromochloromethane was Eastman White Label material. Iodochloromethane was prepared by refluxing sodium iodide with dichloromethane in acetone for 5 days; b.p. 108°, $\lambda_{\text{max}}^{\text{CH}_3\text{CN}} 2630 \text{ \AA}$. ($\epsilon 450$).

Solutions of 1-Ethyl-4-carbomethoxy-pyridinyl in Acetonitrile.—The radical was prepared by zinc metal reduction of the pyridinium iodide in acetonitrile^{2a} and distilled. The distillate was introduced into the apparatus shown in Fig. 6, and diluted with an appropriate amount of acetonitrile. A portion of the solution was poured into the section of the apparatus bearing the quartz cell, both portions of the solution frozen in liquid nitrogen, and the right-hand section sealed off. The concentration of the radical was determined at 6325 Å. using $\epsilon 83$.^{2a} The main body of the solution was poured into six of the breakoff seal tubes, all liquid frozen, and the storage tubes sealed off. The remainder of the solution was introduced into the last six breakoff seal tubes, the liquid frozen, and these were sealed off. The storage tubes were stored in the freezer compartment of a refrigerator. The solutions were a beautiful sapphire blue in color under these conditions.

Radical Stability.—Apparatus A was sealed off below the stopcock (Fig. 7) after the introduction of a solution of the pyridinyl radical 1. The apparatus was thermostated at 50 or at 75° and the loss of radical followed at 6325 Å. At the lower temperature, it was scarcely possible to note any change in concentration. At 75°, a half-life of 12 days for the disappearance of radical was found. This value may be a minimum, but was not further investigated.

(50) Yu. G. Papulov, *J. Structural Chem. USSR*, **4**, 561 (1964).

(51) C. A. Coulson and D. Stocker, *Mol. Phys.*, **2**, 397 (1959).

(52) Reference 12, p. 152.

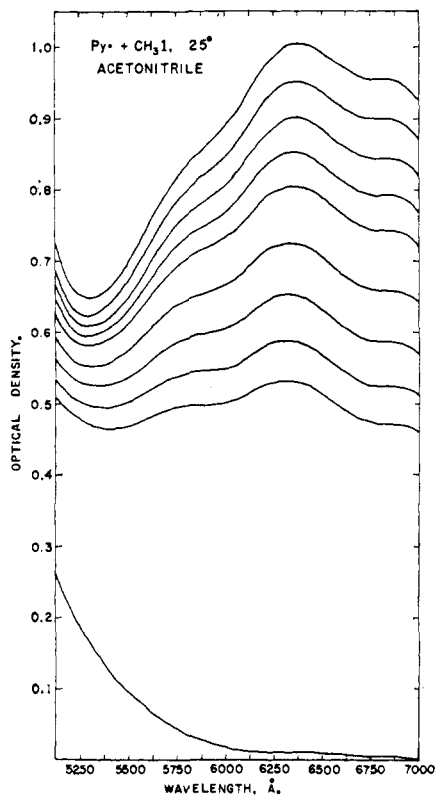


Fig. 8.—Visible absorption curves for the disappearance of 1-ethyl-4-carbomethoxypyridinyl in the presence of excess iodomethane in acetonitrile at 25°.

Rate Measurements.—The basic apparatus (A) for the preparation of the reaction mixtures to be studied is illustrated in Fig. 7. A solution of radical 1 contained in a breakseal storage tube was transferred to tube A and frozen. (In some cases, a major portion of the radical was removed into another breakseal storage tube. This was necessary because radical solutions were usually prepared at a concentration level of *ca.* 0.01 *M*, and for the dilute solutions required in rates measured under bimolecular conditions most of the radical originally added had to be removed.) The halocarbon solution, previously degassed, was distilled into tube B from another vessel on the vacuum line. The solutions were brought to the desired temperature, usually 25°, with a bath, rapidly mixed, the cell filled, and the whole placed in a Cary Model 14 recording spectrophotometer. A black wooden box was used to extend the working space of the cell compartment. In one rate run, it was possible to start recording spectra 55 sec. after mixing; in most cases, from 2 to 3 min. elapsed before recording could begin. The cell was placed in a thermostated cell jacket maintained at the working temperature, but the temperature control was less effective than possible with the bath because the compartment temperatures ranged from 1 to 2° higher. Much better temperature control was obtained when the quartz cell was replaced with a cell made from square Pyrex tubing⁵³ (1-

(53) Precision drawn square Pyrex tubing was obtained from the Kontes Glass Co.

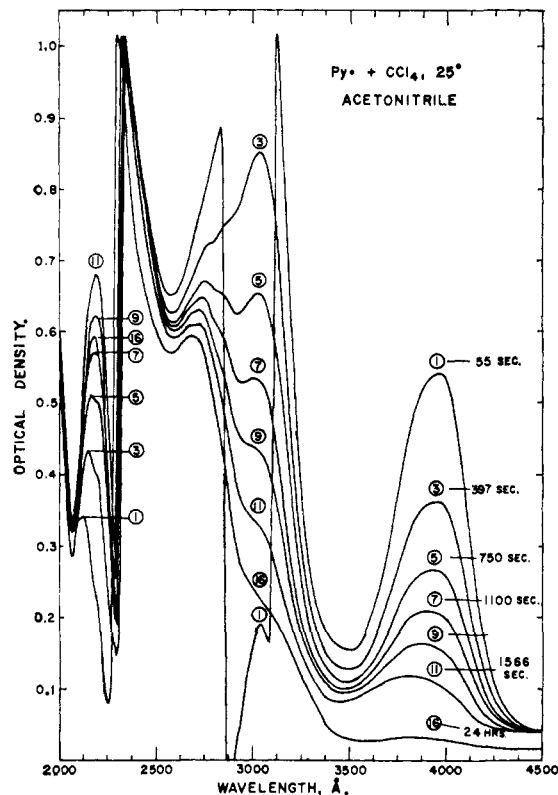


Fig. 9.—Ultraviolet absorption curves showing the change in the concentration of 1-ethyl-4-carbomethoxypyridinyl in the presence of tetrachloromethane in acetonitrile at 25°. Note that only every other curve has been reproduced. The initial concentration of the radical was twice that of the tetrachloromethane.

cm. inside diameter) because the cell could be sealed off from the apparatus A and placed inside the thermostated cell jacket. The disadvantage of the sealed cell technique was that the products could not be examined.

The halocarbon concentration was measured by g.l.c. at the termination of each rate run.

Typical sets of curves are shown in Fig. 8 (iodomethane) and Fig. 9 (tetrachloromethane).

Product Studies.—At the completion of a rate run, the solvent was transferred along with most of the halocarbon to another flask on the line. Hexane was introduced and the hexane-soluble material examined spectroscopically. After transfer of the hexane solution to tube A, water was used to dissolve the residue. The spectrum of the aqueous solution corresponded, as described in the text, to that of 1-ethyl-4-carbomethoxypyridinium ion. The ratio of pyridinium ion found (ϵ 4550 at 2735 Å.) to the pyridinyl radical originally present was: CH_2Br_2 , 0.54; CH_2Br_2 , 0.49; BrCH_2Cl , 0.47; CCl_4 , 0.47; CCl_4 , 0.57; ICH_2Cl , 0.49. A Volhard titration on a micro scale indicated that the ratio of chloride ion to pyridinyl radical was 0.47 ± 0.07 for the case of CCl_4 , in which ($\text{Py}^+/\text{Py}\cdot$) was 0.57.