must be formed in the acid phase before alkylate product is obtained selectively and in close to theoretical yield. Secondly, it has been shown that formation of products other than the octanes proceeds primarily *via* a C$_{12}$-cation intermediate. At low olefin concentration (low space velocity) the C$_{12}^{\oplus}$ comes from alkylation of isobutane with octene. At higher olefin concentration the reaction between octyl cations and butylenes begins to become important. Finally, it has been found that the reactivity of normal and isobutylene under alkylation conditions is considerably different and this leads to vast differences in the formation of alkylate products.

[Contribution from the Department of Chemistry, New York University, New York 53, N. Y.]

Structure and Reactivity in Carbonyl Reactions; Temperature Coefficients of Rate of Formation of Several Thiosemicarbazones[1]

By Irwin D. Fiarman and Joseph D. Gettler

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Rates of thiosemicarbazone formation of seven carbonyl compounds of a variety of structural types have been measured at 25 and 50° in aqueous solution buffered at *p*H 7, using a spectrophotometric method of analysis. Thiosemicarbazone formation for all the compounds investigated is a second-order reaction, and reversible, except for furfural, under the experimental conditions of the investigation. Determination of temperature coefficients of rate permitted the calculation of relative thermodynamic activation quantities. A previously suggested correlation between presumed rigidities of carbonyl compounds and their corresponding activation quantities was not observed. This finding agrees in part with a previous report respecting oxime formation but disagrees with the results established for semicarbazone formation. However, an excellent linear free energy relationship was obtained when the logarithms of the specific rates for thiosemicarbazone formation were plotted against the logarithms of specific rates for semicarbazone formation. The linearity of the free energy relationship suggests the essential validity of the two series of data, the fundamental similarity of the three reaction systems, and the advisability of further study of this particular phase of structure and reactivity. Attempts to obtain linear correlations with Taft σ^*-values failed. In addition, some original spectral data are reported for seven thiosemicarbazones.

Introduction

An especially trenchant analysis of the general problem of structure and reactivity has been presented by Hammett.[2] In addition, Price and Hammett[3] have discussed the effect of structure on reactivity for carbonyl reaction systems and indicated that an explanation of the effect of structural change on reaction rate based exclusively on potential energy effects is valid only where there is a cancellation of other effects due to kinetic energy steric effects. An investigation of temperature coefficients of reaction rate is prerequisite to the determination of the importance of kinetic energy steric effects. Price and Hammett[3] measured the temperature coefficients of the rate of semicarbazone formation and found a parallelism between molecular rigidity and entropy of activation, which is indicated by the data of columns one and two in Table VI. In addition, they proposed that the parallelism would apply with high probability to other typical carbonyl reactions. The parallelism was found subsequent general confirmation in the reaction of benzaldehyde with both acetone and methyl ethyl ketone[4]; in the acid-catalyzed hydration of isobutene, trimethylethylene and 1-methylcyclopentene-1 by aqueous acid[5]; in the hydrolysis of esters catalyzed by ion-exchange resins[6]; in the reaction of methyl iodide with certain amines in nitrobenzene[7,8] and in thiosemicarbazone formation of several alkylbutyrophenones.[9]

In 1956, Fitzpatrick and Gettler,[10] in an attempt to obtain additional support for the Hammett parallelism and to extend the range of its validity, measured the temperature coefficients of rate of oximation. Derived relative entropies of activation are listed in Table VI, column 3, and an examination reveals little or no parallelism between rigidity of structure and entropy of activation.

To resolve this apparent divergence, temperature coefficients of reaction rate of thiosemicarbazone formation were measured, since thiosemicarbazone formation resembles semicarbazone formation more closely, formally or otherwise, than does oximation.

Experimental

Materials.—Thiosemicarbazide hydrochloride (Matheson, Coleman and Bell Inc.) was further purified, stored, and its purity was determined by the method of Price and Hammett.[3] The purified material possessed an acidimetric titer which was within two parts per 1,000 of its iodometric titer.

All carbonyl compounds were distilled in all-glass apparatus before use. Eastman Kodak Co. compounds: acetone was redistilled, b.p. 56.0°; diethyl ketone was redistilled, b.p. 101.5°; methyl ethyl ketone was redistilled, b.p. 79–80°; cyclohexanone was redistilled, b.p. 154–155°; furfural was first redistilled at atmospheric pressure, then further purified by vacuum distillation, b.p. 60–61° at

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[1] Abstract of a dissertation submitted by Irwin S. Fiarman in partial fulfillment of the requirements for the degree of Master of Science in the Faculty of the Graduate School of Arts and Sciences, New York University.

[2] L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, p. 119.

[3] F. P. Price, Jr., and L. P. Hammett, *J. Am. Chem. Soc.*, **63**, 2387 (1941).

[4] J. D. Gettler and L. P. Hammett, *ibid.*, **65**, 1824 (1943).

[5] (a) R. W. Taft, Jr., J. B. Levy, D. Aaron and L. P. Hammett, *ibid.*, **74**, 4735 (1952); **75**, 3955 (1953); (b) E. L. Parlee, R. W. Taft, Jr., and C. A. DeFazio, *ibid.*, **77**, 837 (1955).

[6] (a) V. C. Haskell and L. P. Hammett, *ibid.*, **71**, 1284 (1949); (b) S. A. Bernhard and L. P. Hammett, *bid.*, **75**, 1798 (1953); (c) S. A. Bernhard, E. Garfield and L. P. Hammett, *ibid.*, **76**, 991 (1954); (d) P. Riesz and L. P. Hammett, *ibid.*, **76**, 992 (1954); (e) H. Samelson and L. P. Hammett, *ibid.*, **78**, 524 (1956).

[7] H. C. Brown and A. Cahn, *ibid.*, **77**, 1715 (1955).

[8] H. C. Brown and N. R. Eldred, *ibid.*, **71**, 445 (1949).

[9] J. L. Maxwell, M. Joanne Brownlee and Maretta P. Holden, *ibid.*, **83**, 589 (1961).

[10] F. W. Fitzpatrick and J. D. Gettler, *ibid.*, **78**, 530 (1956).

15 mm., stored under nitrogen at −15°, and used the following day. Matheson, Coleman and Bell Inc. products: cyclopentanone was redistilled, b.p. 131°; pinacolone was redistilled, b.p. 105–106°.

Baker and Adamson sodium phosphates were used to prepare buffer solution.

Samples of pure thiosemicarbazones were prepared according to the method of Shriner and Fuson[11] for semicarbazone preparation. After careful preparation, these substances were used to confirm the validity of the analytical method.

Preparation of Solutions.—Buffer solutions were prepared and their *p*H values confirmed, and thiosemicarbazide solutions and solutions of carbonyl compounds were prepared as previously described.[3]

Apparatus.—For all spectral work a Beckman model DK-2 recording spectrophotometer equipped with a Beckman no. 92430 time drive attachment was used. A pair of matched 1-cm., quartz cells was used for all measurements.

Constant temperatures of 25.00 ± 0.005° or 50.00 ± 0.005° were maintained with an Ultra electronically controlled thermostat.[12]

Method.—Ultraviolet spectrophotometry was employed as a suitable analytical method for the resolution of the problem at hand. An examination of the literature[13] and the results of preliminary experiments revealed that the ultraviolet absorption spectra of thiosemicarbazones differ appreciably from those of thiosemicarbazide and parent carbonyl compounds. Experimentation indicated that this method was suitable to determine the concentrations of either reagent or thiosemicarbazones, with sufficient accuracy to allow the obtention of precise rate and equilibrium constants.

Table I lists the optical absorption characteristics of the substances under investigation.

TABLE I

ABSORPTION MAXIMA IN AQUEOUS SOLUTION OF SEVERAL CARBONYL COMPOUNDS, THEIR CORRESPONDING THIOSEMICARBAZONES AND THIOSEMICARBAZIDE

	Compound		Thiosemicarbazone	
	λ_{max}, mμ	ϵ_{max}	λ_{max}, mμ	ϵ_{max}
Acetone	264	19	264	21,800
Methyl ethyl ketone	266	21	264	20,000
Diethyl ketone	270	25	264	21,700
Pinacolone	273	29	264	23,200
Cyclopentanone	281	24	264	22,500
Cyclohexanone	277	26	264	21,500
Furfural	278	14,800	316	36,800
Thiosemicarbazide	234	11,900	..	....

The feasibility of the spectrophotometric method resides in the following two factors which are readily discernible from an examination of Table I. First, the disappearance of the carbonyl function >C=O and the concomitant formation of a >C=N—linkage results in an appreciable shift in wave length of maximum absorption for most of the compounds listed. Secondly, there exists an extremely large difference in intensity of absorption between thiosemicarbazones (ϵ in excess of 20,000) and carbonyl compounds (ϵ less than 30) in all cases excepting furfural. However, in this latter case the large shift in the wave length of maximum absorption compensates for the very large extinction coefficient of furfural. The operation of these two factors, either separately or in concert, allows obtainment of sufficiently discriminatory analytical data to permit the calculation of precise rate constants. The validity of the conclusions stated immediately above was general for all compounds investigated and successful comprobations of Beer's law were obtained for all the carbonyl compounds under consideration.

Procedure.—The kinetic procedure was essentially that described previously.[10] The method of determination of equilibrium constants was that used previously except that an invariant absorbance over a reasonably long interval of time was accepted as a phenomenological criterion of equilibrium. Since a significant effect of *p*H on rates of oxime formation and semicarbazone formation has been established,[14–16] precautions were taken during the present investigation to establish the constancy of *p*H during the course of individual kinetic experiments.[3] The extent of reagent decomposition during any kinetic experiment was investigated and found to be negligible.

Results

The specific rate data presented in Tables II and III were calculated from the equation previously derived by Conant and Bartlett[14] for semicarbazone formation and used by Price and Hammett.[3] The value of *k* was determined by the application of least squares methods to values of the appropriate concentration function and values of time.

Table II lists detailed data for a kinetic experiment, cyclopentanone thiosemicarbazone formation, while Table III contains a summary of all rate data for thiosemicarbazone formation obtained during the course of the present research. Equilibrium and rate constants for the hydrolysis of several thiosemicarbazones are contained in Table IV.

TABLE II

SAMPLE DATA: CYCLOPENTANONE THIOSEMICARBAZONE FORMATION

Medium: 0.0571 *M* Na_2HPO_4, 0.0286 *M* NaH_2PO_4; temperature 25.00°; initial concn., cyclopentanone 0.04219 *M*; thiosemicarbazide 0.003090 *M*; *k* in l./mole. sec.

Time, sec.	Absorbance	(Reagent)	$({}^1/_2x_e + c + K)\ln(x_e + x_e + c + K)/x - x$	$k \times 10^3$
236	0.033	0.0000680	64.80	..
634	.076	.0001650	65.50	1.76
1321	.137	.0003000	66.59	1.65
1778	.173	.0003870	67.50	1.75
2264	.216	.0004750	68.24	1.69
3080	.281	.0006370	69.75	1.74
3874	.343	.0007620	71.20	1.76
4892	.417	.0009250	72.98	1.75
5241	.438	.0009870	73.68	1.77
5659	.464	.0010280	74.21	1.73
			Av.	1.73 ± 0.03

For all the compounds investigated, the validity of the equation and the accuracy of the results were so satisfactory that least squares analysis showed no trend in the values of the specific rate constants for the entire duration of any one kinetic experiment. Most experiments involved the plotting of a minimum of ten experimental points. The deviation of the mean of the rate constants was taken as the precision of the average since reproducibility was considered a better measure of the accuracy of the rate constant than the probable error in the least squares slope of any individual measurement. This error ranged from 1 to 2%.

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[11] R. L. Shriner and R. C. Fuson, "The Systematic Identification of Organic Compounds," John Wiley and Sons, Inc., New York, N. Y., 1940, pp. 167–168.

[12] Through the courtesy of Dr. Karl Weiss of our staff. For details see ref. 32 in J. Bohning and K. Weiss, *J. Am. Chem. Soc.*, **82**, 4724 (1960).

[13] A. E. Gillam and E. S. Stern, "Electronic Absorption Spectroscopy," Edward Arnold (Publishers) Ltd., London, Eng., 1957, Chapter 5.

[14] J. B. Conant and P. D. Bartlett, *J. Am. Chem. Soc.*, **54**, 2881 (1932).

[15] F. H. Westheimer, *ibid.*, **56**, 1962 (1934).

[16] W. P. Jencks, *ibid.*, **81**, 475 (1959).

Since the problem under consideration is one involving the effect of structural change upon reactivity, the pertinent thermodynamic quantities are not the absolute values of the free energy, $\Delta G^{\ddagger}$, entropy, $\Delta S^{\ddagger}$, and heat, $\Delta H^{\ddagger}$, of activation for a single reaction, but rather the values of these quantities relative to that of a standard reactant, for which acetone has been selected. Such relative quantities were calculated from equations previously reported[2,3] and the results are contained in Table V.

TABLE III

RATE CONSTANTS, k, FOR THIOSEMICARBAZONE FORMATION IN BUFFER OF pH 7.0

Buffer, 0.0571 M Na_2HPO_4, 0.0286 M NaH_2PO_4

Compound	Temp., °C.	[Compound]	[Reagent]	k, l./mole, sec.
Acetone	25.00	0.01665	0.003924	0.00951
		.01959	.004100	.00961
		.01665	.003366	.00953
			Av.	0.00955
	50.00	0.01687	0.003208	0.0125
		.03374	.003452	.0123
		.03504	.003482	.0123
			Av.	0.0124
Diethyl ketone	25.00	0.02252	0.003592	0.00119
		.04050	.003231	.00110
		.02252	.003122	.00113
			Av.	0.00114
	50.00	0.02306	0.003234	0.00219
		.02376	.003326	.00210
		.02306	.003518	.00201
			Av.	0.00210
Pinacolone	25.00	0.03258	0.003910	0.0000622
		.03396	.004068	.0000644
		.03393	.003968	.0000609
			Av.	0.0000625
	50.00	0.03258	0.003910	0.000119
		.03396	.004068	.000114
		.03393	.003968	.000120
			Av.	0.000118
Cyclopentanone	25.00	0.03950	0.002950	0.00176
		.03950	.002742	.00169
		.04219	.003090	.00168
			Av.	0.00171
	50.00	0.03986	0.003086	0.00296
		.03905	.002977	.00307
		.03986	.002916	.00295
			Av.	0.00299
Cyclohexanone	25.00	0.003122	0.001515	0.0528
		.007845	.001515	.0521
		.006509	.001612	.0529
			Av.	0.0526
	50.00	0.005010	0.001961	0.0673
		.004777	.001639	.0666
		.005010	.002232	.0658
			Av.	0.0666
Furfural	25.00	0.04306	0.002906	0.00121
		.04440	.002898	.00120
		.04320	.003182	.00121
			Av.	0.00121
	50.00	0.03940	0.003180	0.00212
		.04350	.003196	.00205
		.03940	.002862	.00212
			Av.	0.00210
Methyl ethyl ketone	25.00	0.02766	0.003754	0.00404
		.02666	.003310	.00418
		.02666	.003014	.00440
			Av.	0.00421
	50.00	0.02666	0.003310	0.00743
		.02666	.003014	.00701
		.02766	.003754	.00745
			Av.	0.00730

TABLE IV

EQUILIBRIUM CONSTANTS, K, AND RATE CONSTANTS, k_h FOR HYDROLYSIS OF SEVERAL THIOSEMICARBAZONES[a]

	Temp., °C.	$K \times 10^3$	$k_h \times 10^4$, sec.$^{-1}$
Acetone	50.00	20.0	2.48
	25.00	7.05	0.672
Diethyl ketone	50.00	102.4	2.15
	25.00	43.6	0.497
Cyclopentanone	50.00	12.3	.368
	25.00	4.08	.0698
Cyclohexanone	50.00	11.3	7.53
	25.00	3.64	1.91
Methyl ethyl ketone	50.00	34.1	2.49
	25.00	9.70	0.408
Pinacolone	50.00	32.7	.0204
	25.00	97.8	.115

[a] Same reaction medium as in Table III.

TABLE V

RELATIVE ENTROPIES, HEATS AND FREE ENERGIES OF ACTIVATION FOR THIOSEMICARBAZONE FORMATION AT 37.5°

Compound	$(\Delta S^{\ddagger} - \Delta S_0^{\ddagger})$, cal./°C.	$(\Delta H^{\ddagger} - \Delta H_0^{\ddagger})$, cal.	$(\Delta G^{\ddagger} - \Delta G_0^{\ddagger})$, alc.
Acetone[a]	0.00	0	0
Diethyl ketone	4.80	2500	1100
Pinacolone	−0.40	2600	2800
Cyclopentanone	4.20	2100	900
Cyclohexanone	2.70	200	−1000
Furfural	8.80	2000	1100
Methyl ethyl ketone	6.40	2000	400

[a] Thermodynamic activation quantities for acetone: $\Delta S_0^{\ddagger}$ −2.80 cal./°C.; $\Delta H_0^{\ddagger}$ 900 cal.; $\Delta G_0^{\ddagger}$ 2800 cal.

Discussion

The spectral data of Table I demonstrate that furfural, while exhibiting maximum absorption at approximately the same wave length as the ketones, possesses an extinction coefficient approximately 500 times greater.

This substantial difference might well be caused by the appreciable amount of conjugation present in the furfural structure. Furthermore, the spectral character of furfural thiosemicarbazone differs appreciably from that of the other six thiosemicarbazones studied. All the ketone thiosemicarbazones show maximal absorption at 264 mμ with an average extinction coefficient of 21,700 ± 900, while furfural thiosemicarbazone exhibits maximal absorption at 316 mμ with an extinction coefficient of 36,800. A shift of such magnitude is typical of aromatic thiosemicarbazones.

The use of an accurate and unambiguous analytical method, the careful experimental design, the care applied in obtaining accurate kinetic data and the failure to fit the resultant data to any reaction order type save that of a reversible second order, all lead to the inescapable conclusion that thiosemicarbazone formation for all compounds studied is unequivocally second order and, save for furfural, reversible, under the experimental conditions employed. It was further observed that for the seven carbonyl compounds examined, under the experimental conditions employed, there was an absence of any appreciable variation in specific rate with variation in the initial reactant concentrations or with time for any one particular kinetic experiment. This suggests the essential soundness of the analytical method and the absence of any environmental factor of appreciable magnitude.

Elements of both similarity and dissimilarity are evident among semicarbazone formation,[3] oxime formation[10] and thiosemicarbazone formation. As one case in point, all these three reactions are kinetically second order. Secondly, all compounds investigated in semicarbazone formation[5] exhibited reversible characteristics. However, furfural and acetophenone possess equilibria so far toward completion that there was no evidence of a trend in the rate constants during the course of the reaction. In thiosemicarbazone formation, furfural is the only compound which exhibits any appreciable amount of irreversibility while oxime formation is irreversible in all cases.

An examination of rate data for semicarbazone formation,[3] oxime formation[10] and thiosemicarbazone formation indicates that for carbonyl compounds, where comparable data exist, rates of oxime formation and thiosemicarbazone formation relative to semicarbazone formation are appreciably greater than unity and less than unity, respectively. Furthermore, the magnitudes of temperature coefficients of rate follow the sequence oximation > semicarbazone formation > thiosemicarbazone formation. These facts indicate that lack of parallelism is not a function of any particular magnitude of rate or of temperature coefficient of a given reaction.

The cardinal objective of the research was to probe further the validity of the Hammett parallelism[3] which is challenged by data on oxime formation.[10] In Table VI are reproduced values of relative activation entropies for the three pertinent reaction systems.

TABLE VI

RELATIVE ENTROPIES OF ACTIVATION FOR SEMICARBAZONE,[3] OXIME[10] AND THIOSEMICARBAZONE FORMATION

	$(\Delta S^{\ddagger} - \Delta S^{\ddagger}_0)$, cal./°C.		
	Semicarbazone formation 12.5°	Oxime formation 12.5°	Thiosemicarbazone formation 37.5°
Cyclopentanone	3.30	0.2	4.20
Furfural	3.20	15.00	8.80
Cyclohexanone	0.40	−5.30	2.70
Acetophenone	−1.90	3.00	..
Acetone	0.00	0.00	0.00
Methyl ethyl ketone	..	4.50	6.40
Diethyl ketone	−6.50	−4.50	4.80
Di-*n*-propyl ketone	..	−15.10	..
Methyl isopropyl ketone	..	6.30	..
Methyl isobutyl ketone	..	7.10	..
Pinacolone	−9.70	7.10	−0.40

An examination of the above three sets of relative entropies shows resemblances and differences. In the case of semicarbazone formation, there is, with the exception of acetophenone, a parallelism between presumed molecular rigidities and corresponding relative entropies of activation. No such parallelism is evident in thiosemicarbazone formation, although there are several points of similarity between the two series of data. The compounds cyclopentanone, furfural, cyclohexanone and pinacolone occupy approximately comparable positions in the two series. The numerical values of the relative activation entropies for these compounds differ in each series, but this is somewhat inconsequential since there is no apparent, theoretical reason to expect identity of magnitudes for compounds in the different reaction series. A notable divergence occurs for methyl ethyl ketone and diethyl ketone which possess positive relative entropies of activation in thiosemicarbazone formation.

A comparison of oxime formation[10] with thiosemicarbazone formation also reveals significant differences. While the compounds cyclopentanone, furfural and methyl ethyl ketone occupy comparable positions in both these series, diethyl ketone, cyclohexanone and pinacolone occupy divergent positions with respect to the sign of relative entropy values. The most serious indictment of the original parallelism[3] resides in the fact that dialkyl ketones exhibit both positive and negative relative entropies of activation of considerable magnitude in both oxime and thiosemicarbazone formation, while the theoretical basis of the parallelism would admit of only negative values.

The relative entropies of activation for thiosemicarbazone formation in the case of the compounds listed in Table VI vary over some nine entropy units, which is equivalent to approximately two orders of magnitude in the reaction rate. It is safe to assume that the variations are many times greater than the probable error in the measurement of the relative activation entropy.

Figure 1 illustrates plots of the logarithms of rate constants for semicarbazone formation[3] *versus* the logarithms of rate constants for thiosemicarbazone formation (line A), and the logarithms of rate constants for oxime formation (line B), taken at 25°. The two plots which represent linear free energy relationships[10,17–19] exhibit a marked parallelism in that line A possesses a slope of 0.93 while

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(17) A. A. Frost and R. G. Pearson, "Kinetics and Mechanism," John Wiley and Sons, Inc., New York, N. Y., 1953, pp. 216–218.

(18) R. L. Hill and T. I. Crowell, Abstracts of Papers, American Chemical Society Meeting, Cincinnati, Ohio, April, 1955, p. 36 N.

(19) Reference 2, pp. 194–238.

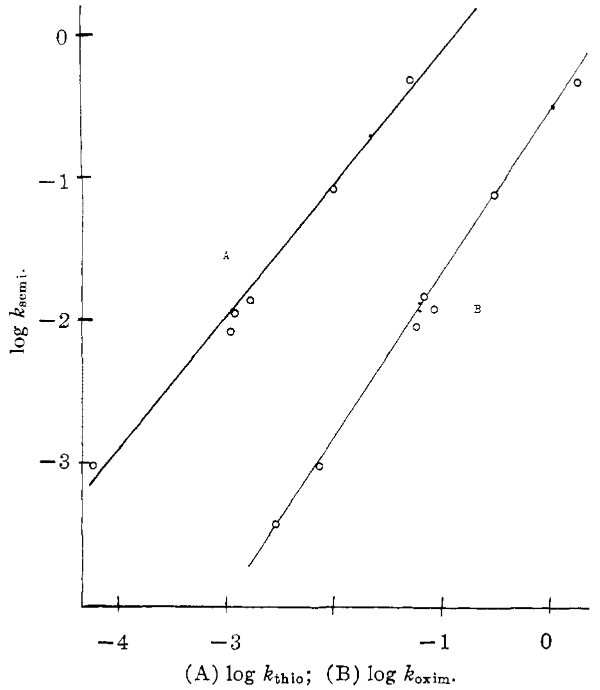

Fig. 1.—Plot of log k for semicarbazone formation *versus* log k for thiosemicarbazone formation (plot A) and log k for oxime formation (plot B). Reading from the uppermost end of the plots downward, the points correspond successively to cyclohexanone, acetone, cyclopentanone, furfural, diethyl ketone, pinacolone and acetophenone.

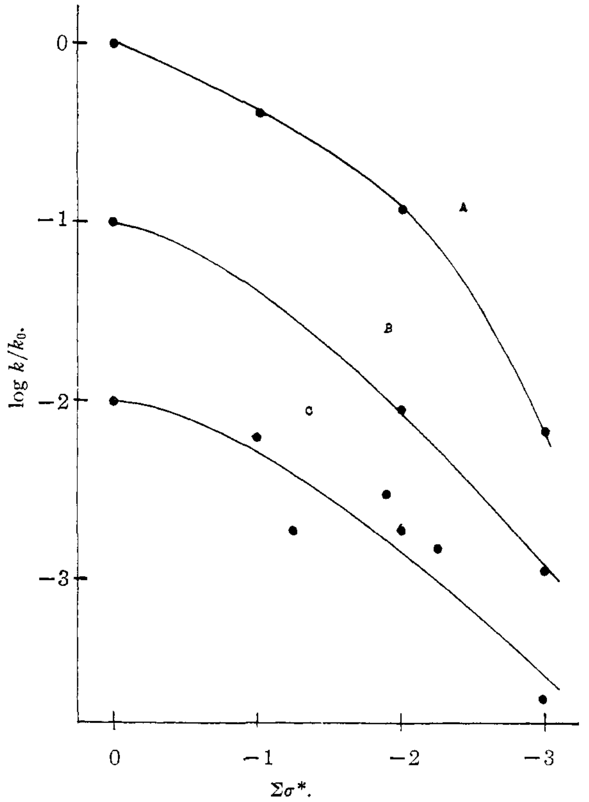

Fig. 2.—Relationships between log (k/k_0) and Taft's σ^* values for (A) thiosemicarbazone formation, (B) semicarbazone formation and (C) oxime formation. From upper left to lower right the points correspond to the compounds: (A) acetone, methyl ethyl ketone, diethyl ketone, pinacolone; (B) acetone, diethyl ketone, pinacolone; (C) acetone, diethyl ketone, methyl isobutyl ketone, methyl isopropyl ketone, diethyl ketone, di-*n*-propyl ketone, pinacolone. The position of the scale of ordinates is arbitrary.

the slope of line B is 1.1 and in the exceedingly high degree of linearity reflected by correlation coefficients of 0.98 and 0.99, respectively.

The high degree of linearity in the free energy plots for the two reaction series under consideration is significant in view of the wide structural variation involved and in that the data of the three systems were obtained by different pairs of investigators employing two fundamentally different experimental techniques. The satisfying linearity can be ascribed to the following three factors. Firstly, it arises, in part, from an implied similarity of the transition states of semicarbazone formation, oxime formation and thiosemicarbazone formation; secondly, from the inherent compressive features of the logarizing process, generally; and, thirdly, from the essential soundness of the experimental data under consideration. The almost perfect linear free energy relationships illustrated in Fig. 1 demands an equally good linear free energy relationship between thiosemicarbazone and oxime formation. It is highly unlikely that this mutuality of relationship is just some mathematical curiosity; it is more probable that the mutuality arises from some inherent similarity in the nature of the three reactions.

Numerous correlations between structure and reactivity have been realized through Hammett's σ–ρ function,[20] and Taft[21] has extended the use of this relationship to aliphatic compounds by calculating σ^*-values from the rates of hydrolysis of esters. In many cases,[22] a wide variety of reaction types fit with relatively good precision an equation, analogous to that of Hammett with respect to the proportionality of polar effects, namely

$$\log (k/k_0) = \sigma^*\rho^* \qquad (1)$$

where σ^* is the polar substituent constant for the group R relative to the standard CH_3 group; ρ^* is a constant giving the susceptibility of a given reaction series to polar substitution.

Figure 2 demonstrates the result of plotting Taft σ^*-values against corresponding log (k/k_0) values for thiosemicarbazone formation (curve A), semicarbazone formation (curve B) and oxime formation (curve C). The resultant plots are decidedly nonlinear in character. The realization of these nonlinear relationships does not necessarily contraindicate the proportional nature of polar effects, but shows that in these three cases rates are probably determined by other factors. The approximately mutual parallelism of the curves indicates some common, internal factor of resemblance.

The results of this research strongly suggest that

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(20) (a) L. P. Hammett, *Chem. Revs.*, **17**, 125 (1935); (b) "Physical Organic Chemistry," McGraw–Hill Book Co., Inc., New York, N. Y., 1940, p. 184.

(21) R. W. Taft, *J. Am. Chem. Soc.*, **74**, 3120 (1952); **75**, 4231 (1953).

(22) M. Newman, "Steric Effects in Organic Chemistry," John Wiley and Sons, Inc., New York, N. Y., 1956, p. 638.

an explanation for the failure of the Hammett parallelism in the present case might reside in an investigation of the temperature coefficients of rate of the addition and dehydration sub-steps postulated by Jencks[16] for semicarbazone formation and oxime formation.

[CONTRIBUTION NO. 708 FROM THE CENTRAL RESEARCH DEPARTMENT, EXPERIMENTAL STATION, E. I. DU PONT DE NEMOURS AND CO., WILMINGTON 98, DEL.]

The Chemistry of Diazaquinones. 3,6-Pyridazinedione and 1,4-Phthalazinedione

BY THOMAS J. KEALY

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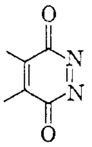
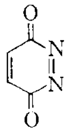
The unstable diazaquinones 3,6-pyridazinedione, 4,5-difluoro-3,6-pyridazinedione and 1,4-phthalazinedione have been prepared by the low temperature oxidation of alkali metal salts of the corresponding hydrazides using *t*-butyl hypochlorite. The decomposition of the diazaquinones has been investigated. The diazaquinones react instantly with dienes at −77° and therefore are among the most reactive of the dienophiles. Use of the diazaquinones in the Diels–Alder reaction has made some otherwise difficultly accessible ring systems readily available.

Cyclic azo compounds have been used to advantage in the synthesis of ring compounds of various sizes by the elimination of nitrogen.[1,2] Some diacyl diimides (RCON=NCOR) and diaroyl diimides afford 1,2-diketones as the principal product when decomposed under certain conditions.[3,4] However, relatively little is known of cyclic systems containing the diacyl diimide grouping. Of particular interest to us was the preparation of compounds of type I since azo compounds in general afford radicals on decomposition[5] and those one might expect from compounds of this type seemed especially appealing.

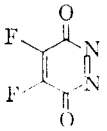


I

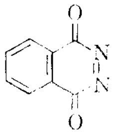
We chose as the route to such compounds the low-temperature oxidation of the corresponding cyclic hydrazides. *t*-Butyl hypochlorite was found to be particularly effective for this oxidation since it is soluble in various organic solvents at low temperatures and does not give large amounts of inorganic products. Using this reagent with alkali metal salts of the corresponding hydrazides we have prepared 3,6-pyridazinedione (II), 4,5-difluoro-3,6-pyridazinedione (III) and 1,4-phthalazinedione (IV), for which we propose the generic term, diazaquinones. Clement[6] has recently reported the preparation of unstable acetonitrile solutions of IV by the oxidation of phthalhydrazide at 0° using lead tetraacetate, and has characterized the butadiene adduct of IV.



II

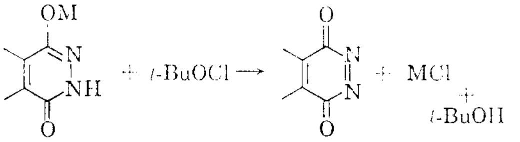


III



IV

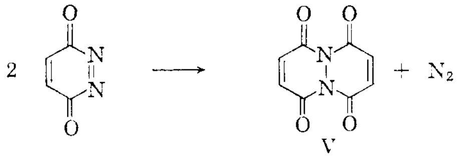
The preparations were conveniently carried out in acetone at −50 to −77°, according to eq. 1, to obtain emerald-green solutions of the diazaquinones. The insoluble salts were removed by filtration at −77° under nitrogen. The diazaquinones are very unstable and highly reactive, and, although both II and IV are precipitated from acetone solutions upon the addition of highly chlorinated solvents, only IV has been isolated. No attempts were made to isolate III. Although IV has been isolated as a crystalline green solid, it frequently decomposed suddenly during the isolation procedures. In acetone solution at −77°, however, the diazaquinones are stable at least overnight and probably longer.



(1)

Decomposition of Diazaquinones.—Slow warming of diazaquinone solutions to room temperature results in a gradual fading of the green color well below 0°. 3,6-Pyridazinedione decomposes rapidly in solution near −30°, and the difluoro derivative III decomposes rapidly above −20°, the solution becoming colorless by 0°. 1,4-Phthalazinedione, which appears to be the most stable of the three, retains only a pale lime-green color on reaching room temperature in solution.

In the case of 3,6-pyridazinedione, warming of the solution results in nitrogen evolution and the formation of 1,4,6,9-tetraketopyridazino(1,2-a)-pyridazine (V) as the principal product (43% yield). This yellow compound was identified by hydrolysis



V

(2)

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[1] For a review, see C. G. Overberger, *Record Chem. Progr.*, **21**, 21 (1960).

[2] R. Criegee and A. Rimmelin, *Chem. Ber.*, **90**, 414 (1957).

[3] H. H. Inhoffen, H. Pommer and F. Bohlmann, *ibid.*, **81**, 507 (1948).

[4] L. Horner and W. Naumann, *Ann.*, **587**, 93 (1954).

[5] On the question of the formation of free acyl and aroyl radicals by decomposition of the corresponding diimides, see ref. 4 and R. Cramer, *J. Am. Chem. Soc.*, **79**, 6215 (1957).

[6] R. A. Clement, *J. Org. Chem.*, **25**, 1724 (1960).