solid was subjected to the following hydrolysis experiments.

(a) A mixture of 1.00 g. of the solid and 50 cc. of distilled water was shaken for 5 minutes. The resulting mixture, consisting of a black tar and red solution, was extracted thoroughly with ether. The ether solution was washed with 6 *N* hydrochloric acid, dried over Drierite and the ether distilled. The residual yellow oil gave 0.33 g. of benzaldehyde 2,4-dinitrophenylhydrazone when treated in the usual way.

The aqueous layer was combined with the acid wash and made basic with sodium bicarbonate. Quinaldamide (III), 0.43 g., was collected by filtration and an additional 0.03 g. was obtained by ether extraction of the alkaline filtrate. No copper quinaldate was isolated after acidification of the alkaline solution with acetic acid and treatment with copper sulfate.

(b) About 1.5 g. of the solid was washed thoroughly with anhydrous benzene to remove any adhering benzaldehyde. The insoluble residue weighed 1.32 g. after having been dried in a vacuum desiccator. This was added to 50 cc. of 5% sodium carbonate solution and shaken for 8 minutes. There was an immediate reaction to give a yellow solid and a black tar The mixture was worked up as described under (a), and there was obtained 0.32 g. of benzaldehyde 2,4-dinitrophenylhydrazone and 0.31 g. of quinaldamide (III). A few mg. of unidentified material of m.p. 220–221° was also isolated.

Action of Hydrogen Chloride on 1-Benzoyl-1,2-dihydroquinaldonitrile (V) in Benzaldehyde Solution.—Anhydrous hydrogen chloride was bubbled into a suspension of 10.00 g. of V in 50 cc. of freshly distilled benzaldehyde for 15 minutes and then the reaction mixture was allowed to stand overnight. The mixture was filtered to remove ammonium chloride which had precipitated, and 12 cc. of concentrated hydrochloric acid was added to the filtrate. More ammonium chloride precipitated and was removed by filtration. Addition of 100 cc. of ether and 20 cc. of ethanol to the filtrate caused 2.48 g. of an orange solid A to precipitate. The mother liquor was extracted with dilute hydrochloric acid. Upon making the acid extract basic by addition of sodium bicarbonate, 0.32 g. of an unidentified yellow solid precipitated, m.p. 100–110° dec. Extraction of the alkaline filtrate with ether afforded a trace of organic material. After acidification of the alkaline filtrate with acetic acid, 1.75 g. of quinaldic acid was obtained *via* the copper salt.

The benzaldehyde–ether layer was washed with sodium bicarbonate solution, the ether distilled and the benzaldehyde steam distilled. Upon trituration of the gummy, yellow residue with a small amount of ethanol, 3.44 g. of colorless solid, m.p. 100–146°, was obtained. This consisted of a mixture of benzoin quinaldate (IV) and 2-(2-quinolyl)-4,5-diphenyloxazole (XXIII). By fractional crystallization from ethanol there was obtained 0.44 g. of pure benzoin quinaldate (IV), m.p. 163–164°, also in admixture with an authentic sample of IV, as the least soluble fraction. From the mother liquors there was obtained 0.28 g. of pure 2-(2-quinolyl)-4,5-diphenyloxazole (XXIII), m.p. 140.5–141.5° after recrystallization from ligroin.

Anal. Calcd. for $C_{24}H_{16}N_2O$: C, 82.74; H, 4.63; N, 8.04. Found: C, 82.87; H, 5.08; N, 8.29.

To 2.00 g. of the orange solid A was added 50 cc. of distilled water and the mixture was stirred for 30 minutes. A precipitate which formed weighed 0.23 g. and consisted of a mixture of IV and XXIII. The aqueous filtrate was extracted with ether. Evaporation of the ether gave no benzaldehyde and only a trace of organic material. The aqueous residue was made basic by addition of sodium bicarbonate and extracted with ether. Again, distillation of the ether left only a trace of organic material. The aqueous solution was acidified with acetic acid, and 0.42 g. of quinaldic acid (II) was isolated *via* the copper salt.

Acid-catalyzed Hydrolysis of 2-(2-Quinolyl)-4,5-diphenyloxazole (XXIII).—A mixture of 0.25 g. of XXIII and 10 cc. of 25% sulfuric acid was heated under reflux for 6 days. Some solid which had steam distilled into the condenser was washed down with water. The mixture was extracted with ether, and the ether solution was washed with sodium bicarbonate solution. No organic material was obtained upon acidification of the sodium bicarbonate solution. The ether solution was dried over anhydrous sodium sulfate and the ether distilled. There remained 0.07 g. (89% net yield) of benzoin, m.p. 132–133° after two crystallizations from ethanol, also in admixture with authentic benzoin.

The original acid layer was made basic by addition of sodium hydroxide solution. Ether extraction gave 0.12 g. of recovered XXIII.

After acidification of the basic solution with acetic acid, a few mg. of quinaldic acid was isolated *via* the copper salt.

LAWRENCE, KANSAS

[CONTRIBUTION FROM THE DEPARTMENT OF CHEMISTRY, DUQUESNE UNIVERSITY]

The Basicity of Hydrazones[1]

BY HUGH F. HARNSBERGER, EDWARD L. COCHRAN AND H. HARRY SZMANT[2]

RECEIVED APRIL 6, 1955

Acidity constants for a number of substituted acetophenone and benzophenone hydrazones were determined spectrophotometrically. The resultant *pK* values can be correlated with the Hammett σ-values when the hypothetical formaldehyde hydrazone is considered to be the parent structure and when the sum of the σ-constants of all substituents is employed.

Introduction

The relative basicity of hydrazones is probably of some significance in the acid-catalyzed conversion of hydrazones to azines[3] since the reaction proceeds with increasing difficulty as one compares dialkyl, alkaryl and diaryl ketone hydrazones. Also, in the case of the Wolff–Kishner reaction it was shown[4] that the base-dependent reaction involves the formation of the hydrazone anion, and while it is experimentally difficult to determine the relative acidity of various hydrazones it is to be expected that the relative acidities are inversely related to the basicities studied here.

Experimental

Since hydrazones tend to hydrolyze rather readily under acidic conditions all of the work reported here was performed using an essentially anhydrous methanol medium. Known volumes of standard solutions of trichloroacetic acid and of the hydrazone were mixed and the resulting hydrogen ion concentration was immediately measured by the spectrophotometric method of Brode.[5] This method requires the measurement of the extinction coefficient of a suitable indicator at its wave length of maximum absorption in a series of buffer solutions covering the useful *p*H range of the indicator. The *p*H of any solution can then be determined by reference to the plot of the extinction coefficient *vs.* *p*H.

Trichloroacetate buffers were prepared by adding known volumes of standard trichloroacetic acid solutions to stand-

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(1) From the M.S. thesis of E. L. C., Duquesne University, 1951.

(2) To whom requests for reprints should be sent.

(3) H. H. Szmant and C. McGinnis, THIS JOURNAL, **72**, 2890 (1950).

(4) H. H. Szmant, H. F. Harnsberger, T. J. Butler and W. P. Barie, *ibid.*, **74**, 2724 (1952).

(5) W. R. Brode, *ibid.*, **46**, 581 (1924).

ard sodium methoxide solutions, and the buffer compositions were calculated so that an ionic strength of 0.0010 was maintained in every case. The ionization constant of trichloroacetic acid in methanol was recalculated[6] from the conductance data of Goldschmidt[7] for trichloroacetic acid, hydrochloric acid, sodium chloride and sodium trichloroacetate. The following values of the ionization constants of trichloroacetic acid and thymol blue in methanol were employed:

Acid	Ionic strength	$K_a \times 10^5$ (25°)
Trichloroacetic	0	1.42
Trichloroacetic	0.0010	1.89
Thymol blue	.0010	2.9

The methanol employed in this work was prepared from Mallinckrodt AR grade by distillation from magnesium turnings. Sodium methoxide solutions were prepared by allowing metallic sodium to react with the above redistilled methanol, and were standardized with hydrochloric acid using brom cresol green indicator. Trichloroacetic acid (C.P.) was vacuum distilled before use, and thymol blue (Eastman Kodak Co.) was used without further purification. Acetophenone hydrazone and its derivatives were prepared by the method of Lock and Stach.[8] *p*-Methoxyacetophenone hydrazone[9] was not found described in the literature. The other hydrazones were prepared previously in this Laboratory.[3] All the hydrazones were either recrystallized to constant melting point or vacuum distilled before use.

Solutions of trichloroacetic acid and sodium methoxide of suitable concentrations were added to 10-ml. volumetric flasks from semi-micro burets in an amount calculated to give the desired *p*H and an ionic strength of 0.0010. One ml. of thymol blue solution (approximately 100 mg./l.) was then added and the solution diluted to the mark. The extinction coefficient of the indicator was measured using a Beckman DU spectrophotometer and a plot of the extinction coefficient *vs.* *p*H was made over the useful range of the indicator, namely, 3.2–5.5 *p*H.

For a given measurement at a desired initial acid concentration the *p*H was calculated for ionic strength 0.0010 in the following manner. By measuring the extinction coefficients after successively larger volumes of hydrazone solution of suitable concentration had been added, it was possible to obtain by interpolation the volume required to give the desired extinction coefficient (and thus *p*H) at fixed ionic strength. The concentration of hydrazonium ion (BH^+) was taken as the difference between 0.0010 *M* (the trichloroacetate ion concentration) and the hydrogen ion concentration. The concentration of the hydrazone (B) was taken as the difference between total added hydrazone (B_0) and the hydrazonium ion concentration. The acidity constant for each hydrazone was calculated for the ionization equilibrium of BH^+.

$$BH^+ \rightleftarrows H^+ + B \qquad K_a = (H^+)(B)/(BH^+)$$

An example is given to make the method of calculation clear.

For acetophenone hydrazone (APH): using $(CCl_3COOH)_0 = 0.00314\ M = (HA)_0$; when $(CCl_3CO_2^-) = 0.00100\ M$

$$pH = pK_{HA} - \log\{[(HA)_0 - (A^-)]/(A^-)\} = $$
$$4.72 - \log[(0.00314 - 0.00100)/(0.00100)] = 4.39$$

At this *p*H the extinction coefficient of thymol blue (from calibration curve)

$$\epsilon = 44.7\ \text{l. g.}^{-1}\ \text{cm.}^{-1}$$

Ml. of 0.00754 *M* APH soln. added to a tot. vol. of 10 ml.	0.55	1.00	1.50	2.00	2.50
Measd. ϵ, l. g.$^{-1}$ cm.$^{-1}$	61.5	55.8	49.3	43.4	39.1

Thus volume of 0.00754 *M* APH to obtain ϵ = 44.7 is 1.88 ml./10 ml. total

$$(BH^+) = (A^-) - (H^+) = 0.00100 - 0.00004 = 9.60 \times 10^{-4}$$
$$(B) = (B)_0 - (BH^+) =$$
$$(1.88)(7.54 \times 10^{-3}) - 9.60 \times 10^{-4} = 4.58 \times 10^{-4}$$

Hence $K_a = 1.93 \times 10^{-5}$ and $pK_a = 4.71$

This procedure was repeated at several initial acid concentrations with solutions made up from different samples of hydrazone. Acid and indicator solutions were used the same day they were prepared in order to minimize the effect of esterification.

Since a thermostated cell holder for the spectrophotometer was not available during the early part of this study a careful determination of the temperature dependence of the *pK* of benzophenone hydrazone was made. The results listed in Table I show that the temperature coefficient is rather small (approximately −0.012 *pK* unit/°C.). Thus the deviations in the *pK* values which were obtained at 22 ± 3° (Table II) can be attributed to the experimental error inherent to the method rather than being caused by the temperature fluctuations.

TABLE I

THE TEMPERATURE DEPENDENCE OF THE pK_a OF BENZOPHENONE HYDRAZONE

Temp., °C.	$pK_{BPH_2^+}$	*p*H	(TCA)/(A⁻)	$(BPH)_0 \times 10^3$	pK_{TCA}	(BPH)/(BPH_2^+)
27.6	3.89	4.72	0.95	7.6	4.70	6.8
27.6	3.86	4.40	2.0	4.3	4.70	3.5
28.0	3.85	4.40	1.99	4.54	4.70	3.6
28.0	3.87	4.72	0.95	7.85	4.70	7.0
35.0	3.74	4.37	1.95	5.03	4.66	4.3
35.0	3.74	4.07	3.90	2.86	4.66	2.1

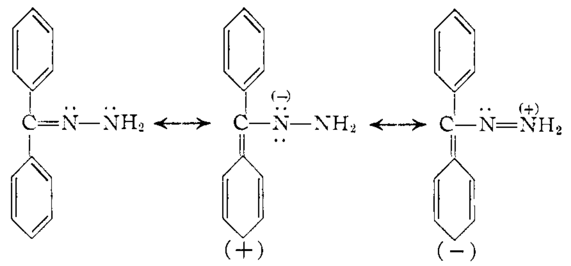
TABLE II

THE pK_a VALUES OF VARIOUS HYDRAZONES (AT 22 ± 3°)

Hydrazone of	Av. pK_a	No. of determn.
Benzophenone	3.85 ± 0.03	7
p,p'-Dimethoxybenzophenone	4.38 ± .05	6
p,p'-Dichlorobenzophenone	3.13 ± .03	6
p-Chlorobenzophenone	3.53 ± .03	6
Phenyl 2-thienyl ketone	3.80 ± .04	6
Acetophenone	4.70 ± .04	5
p-Methoxyacetophenone	4.94 ± .05	5

Discussion

The consideration of the resonance structures of benzophenone hydrazone shows that the polarization of the molecule can occur in either way as to give an electron accumulation at the nitrogen group or in the benzene ring.[10]



The accumulation of the negative charge at the nitrogen terminus of the hydrazone molecule should increase the basicity of the compound and the contribution of such resonance structure is large when the benzene ring carries electron-releasing substit-

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(6) See D. A. McInnes, "The Principles of Electrochemistry," Reinhold Publishing Corp., New York, N. Y., 1939, p. 363.

(7) H. Goldschmidt, *et al.*, *Z. physik. Chem.*, **108**, 123 (1921); **114**, 1 (1924); **117**, 316 (1925).

(8) G. Lock and K. Stach, *Ber.*, **77B**, 293 (1944).

(9) A white crystalline solid, m. 135°. *Anal.* Calcd. for $C_9H_{12}N_2O$: C, 65.83; H, 7.36. Found: C, 65.87; H, 7.63.

(10) H. H. Szmant and C. McGinnis, THIS JOURNAL, **74**, 240 (1952).

uents such as the methoxy group. Conversely, the presence of electron-withdrawing substituents such as the chloro group, should lower the basicity below that of the parent hydrazone, and this effect is indeed observed (Table II). Noteworthy is the consistent difference between the diaryl and methyl aryl ketone hydrazones. Apparently the greater electron-releasing character of the methyl group is more significant in affecting the relative basicities of the hydrazones than is the greater opportunity for resonance stabilization when an aryl group is present.

In spite of the limited number of compounds studied here it is of interest to report certain interesting correlations of the observed pK values with the σ-constants of Hammett.[11]

In his recent review of the Hammett equation Jaffe[12] discusses the additivity of σ-constants when reactions involving compounds containing two substituted benzene rings are compared. He also suggests the possibility of the application of the σ-constants in reaction series of compounds other than aromatic ring systems provided that the rigidity of the molecules assures a constancy of the entropy term.

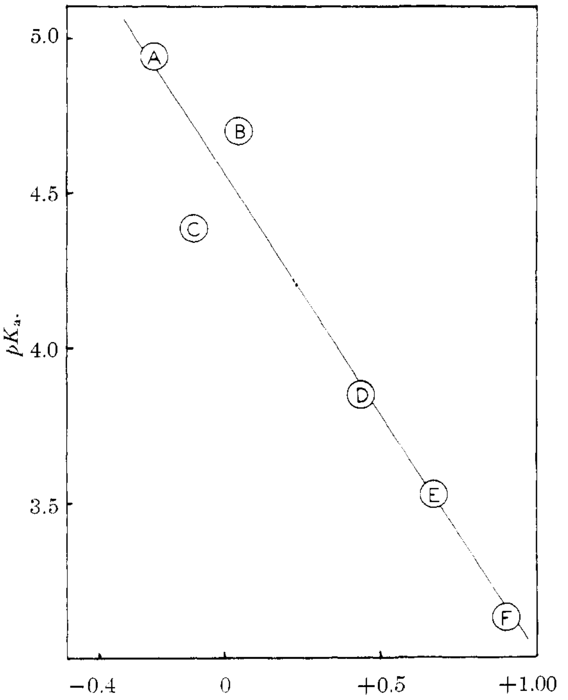

Fig. 1.—Hydrazone acidity constants *vs.* sum of substituent σ values: A, *p*-methoxyacetophenone; B, acetophenone; C, *p,p*-dimethoxybenzophenone; D, benzophenone; E, *p*-chlorobenzophenone; F, *p,p'*-dichlorobenzophenone.

It was observed that while the sum of the σ-constants of the substituted benzophenone hydrazones when plotted against the observed pK values gave points which fitted rather well a straight line, the points for the acetophenone hydrazones showed considerable deviation from this linear plot. The pK values of all the hydrazones could be placed on a single straight line (Fig. 1) when the hypothetical formaldehyde hydrazone was assumed to be the parent compound and the replacement of each of the hydrogens of the formaldehyde residue by methyl, phenyl or substituted phenyl groups was taken into account in the algebraic sum of the corresponding σ-values. The *para*-σ values as given by Hammett[11] were used except for the value for the *p*-phenyl group (+0.22).[13] The calculation of the sum of the σ-constants is illustrated for the case of *p*-methoxyacetophenone hydrazone

$$\Sigma\sigma = \sigma_{p\text{-phenyl}} + \sigma_{p\text{-methoxy}} + \sigma_{p\text{-methoxy}} = 0.22 - 0.27 - 0.17 = -0.22$$

If the above treatment is assumed to be valid then it permits estimating the basicities of hydrazones which are too unstable for experimental study. Thus, the basicity of acetone hydrazone is estimated to have a pK 5.1. Also, from the observed pK of phenyl 2-thienyl ketone hydrazone it is estimated that the σ-constant of the 2-thienyl group has the value of +0.27. The slightly larger positive value of this σ-constant as compared to the value used here for the *p*-phenyl group (+0.22) is in agreement with the greater acidity of 2-thiophenecarboxylic acid as compared with benzoic acid.[14]

The above extension of the application of the Hammett equation is not unreasonable when one considers that in $RR'C{=}N{-}NH_2$ the substituents R and R' are separated from the terminal amino group by a rigid and readily polarizable carbon-nitrogen double bond. It should be pointed out that the entropy effects in the reaction studied here are fortunately rather constant because the steric effects would be expected to be insignificant, in the transition state $RR'C{=}N{-}NH_2{}^{-\delta} \ldots H \ldots {}^{-\delta}O_2C{-}CCl_3$. The same may not be true if the reaction of the hydrazones were one involving a more crowded transition state, such as, for example, the conversion of the hydrazones to azines.[3]

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[11] L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, p. 188.

[12] H. H. Jaffe, *Chem. Revs.*, **53**, 101 (1953).

[13] This is the value suggested by E. S. Gould and J. D. McCullough, This Journal, **73**, 1109 (1951). The variability of the σ constant for the *p*-phenyl substituent was discussed recently by E. Berliner and Liang Huang Lin, *ibid.*, **75**, 2417 (1953).

[14] W. Catlin, *Iowa State Coll. J. Sci.*, **10**, 65 (1935), reports the ionization constant for 2-thiophenecarboxylic acid as 24×10^{-5}. Thus this acid is approximately five times stronger than benzoic acid.