Thermodynamic Data for the Formation of Molecular Complexes between Phenyl Substituted Amides and Iodine

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The free energies and enthalpies of association of the addition compounds formed between various phenyl substituted amides and iodine are reported. The data are correlated through Hammett's equation and an "*ortho* effect" is detected. Structures are discussed. Enthalpies are determined by two different methods and the temperature dependence of the molar absorptivities of the complexes is observed. Support is gained for determining enthalpies of association by a previously reported spectrophotometric method in which certain wave lengths are selected for measurements where temperature independent molar absorptivities are indicated.

Introduction

In recent publications[2,3] a procedure is reported to enable more accurate spectrophotometric determination of enthalpy values for the formation of molecular addition compounds of iodine. In this procedure the enthalpy is determined from a single solution eliminating inaccuracies introduced into the standard procedure by concentration errors. Careful examination of the entire spectrum at several temperatures is necessary to enable selection of a wave length for measurements where the molar absorptivity of the complex does not change appreciably with temperature variation.[3]

The accuracy gained by employing this procedure is demonstrated by the values obtained for the enthalpy of adduct formation between iodine and a series of phenyl substituted amides. If the standard procedure for enthalpy determination were employed for this series, the differences in the results obtained for most of these amides would be inside experimental error at the confidence level employed. Differences could be discerned for many of these compounds by our procedure. The spectral characteristics of the iodine–amide systems studied here are found to parallel closely those of the dimethylbenzamide–iodine system, which has been studied extensively.[3] This makes possible the accurate determination of heats of formation at the same wave length for all systems reported in this paper. The relative order indicated by the measured heats of formation is found to be independent of the wave length at which they are measured. The absolute values of ΔH are obtained at a wave length believed to be at or near the wave length at which molar absorptivities do not change with temperature. These values are compared with those obtained by a more laborious procedure in which the changes in molar absorptivity of the complex with temperature are measured and corrections are made for the change in calculating the enthalpy.

The results from this study are of additional interest because they afford an opportunity to correlate thermodynamic data for the formation of iodine adducts in a non-polar solvent with Hammett's σ values.[4]

Experimental

General Procedures.—The procedures for carrying out the spectrophotometric measurements and calculating the equilibrium constants and enthalpy values have been described.[3] In all experiments $A - A^0$ values (*i.e.*, absorbance of the solution minus that of the initial iodine solution) were read directly by using iodine in CCl₄ as the reference and CCl₄ solution of amide plus iodine in the sample cell. The same total concentration of iodine was employed in the reference and sample.

Error Analysis.—Enthalpy values were determined and the error calculated by the method of least squares. Errors in the

(1) General Electric Foundation Predoctoral Fellow, 1961–1962. Abstracted in part from the Ph.D. thesis of R. L. Carlson, Univ. of Illinois, Urbana, Illinois, 1962.

(2) R. S. Drago, R. L. Carlson, N. J. Rose and D. A. Wenz, *J. Am. Chem. Soc.*, **83**, 3572 (1961).

(3) R. L. Carlson and R. S. Drago, *ibid.*, **84**, 2320 (1962).

(4) M. S. Newman, "Steric Effects in Organic Chemistry," John Wiley and Sons, Inc., New York, N. Y., 1956, p. 571.

equilibrium constant are reported at the 90% confidence level. In an enthalpy experiment, in which a constant $\epsilon_C - \epsilon_I$ (molar absorptivities of the complex minus the iodine) is assumed, the data fit the least squares line extremely well and the error is often less than 0.1 kcal. mole⁻¹. Different investigators obtained results that agree to better than 0.1 kcal. mole⁻¹. The principal source of error in this kind of experiment is caused by the change in $\epsilon_C - \epsilon_I$ with temperature. The reported error in the enthalpy is obtained by estimating the change in the quantity $\epsilon_C - \epsilon_I$ and calculating what effect this has on $-\Delta H$.

Preparation and Purification of Materials.—Baker and Adamson iodine was resublimed three times and stored in a desiccator over phosphorus pentoxide.

Fisher Spectroanalyzed carbon tetrachloride was employed as a solvent in all measurements made.

The substituted N,N-dimethylbenzamides were prepared from the respective substituted benzoic acids. The acid was refluxed with excess thionyl chloride until the evolution of hydrogen chloride ceased. The excess thionyl chloride was removed by distillation at atmospheric pressure and the substituted benzoyl chloride was obtained by distillation at reduced pressure. The freshly distilled acid chloride was dissolved in ether and added dropwise, with stirring, to an ether solution saturated with dimethylamine at 0°. Following the addition of the acid chloride excess dimethylamine was bubbled through the solution. The ammonium salt was then filtered off, the ether solution extracted with water and dried over anhydrous magnesium sulfate. The ether was distilled off to give a mixture containing the desired amide. The monosubstituted *ortho* derivatives are liquids at room temperature and were purified by distillation through a column packed with glass helices at reduced pressures. The remaining amides were obtained pure by crystallization and recrystallization from an ether–petroleum ether mixture. The elemental analyses obtained on the substituted N,N-dimethylbenzamides are:

p-Chloro-N,N-dimethylbenzamide.—*Anal.* Calcd. for C₉H₁₀ClNO: C, 58.86; H, 5.47; N, 7.63. Found: C, 58.63; H, 5.50; N, 7.38.

o-Chloro-N,N-dimethylbenzamide.—*Anal.* Calcd. for C₉H₁₀ClNO: C, 58.86; H, 5.47, N, 7.63. Found: C, 58.61; H, 5.43; N, 7.63.

3,4-Dichloro-N,N-dimethylbenzamide.—*Anal.* Calcd. for C₉H₉Cl₂NO: C, 49.57; H, 4.16; N, 6.42. Found: C, 49.77; H, 4.26; N, 6.22.

2,4-Dichloro-N,N-dimethylbenzamide.—*Anal.* Calcd. for C₉H₉Cl₂NO: C, 49.57; H, 4.16; N, 6.42. Found: C, 49.87; H, 4.33; N, 6.37.

p-Methyl-N,N-dimethylbenzamide.—*Anal.* Calcd. for C₁₀H₁₃NO: C, 73.59; H, 8.03; N, 8.59. Found: C, 73.36; H, 8.06; N, 8.45.

o-Methyl-N,N-dimethylbenzamide.—*Anal.* Calcd. for C₁₀H₁₃NO: C, 73.59; H, 8.03; N, 8.59. Found: C, 73.63; H, 7.89; N, 8.53.

p-Methoxy-N,N-dimethylbenzamide.—*Anal.* Calcd. for C₁₀H₁₃NO₂: C, 67.02; H, 7.31; N, 7.82. Found: C, 67.15; H, 7.30; N, 7.99.

o-Methoxy-N,N-dimethylbenzamide.—*Anal.* Calcd. for C₁₀H₁₃NO₂: C, 67.02; H, 7.31; N, 7.82. Found: C, 66.83; H, 7.23; N, 7.80.

Results

The equilibrium constants and the differences in molar absorptivities between the complexes and iodine are reported in Table I. C_D refers to the initial concentration of donor and C_I to the initial iodine concentration. All measurements were made at 25.0 ± 0.1°.

The results of enthalpy experiments assuming a temperature independent $\epsilon_C - \epsilon_I$ are reported in Table II. As previously pointed out[3] the best enthalpy values for these systems are those calculated near the complex

TABLE I
SPECTROPHOTOMETRIC DATA EMPLOYED FOR EQUILIBRIUM CONSTANT[a] CALCULATIONS

Amide	Wave length, mμ	$C_I \times 10^3$	$C_D \times 10^2$	$-(A - A^0)$
p-Chloro-N,N-di-			7.26	0.148
methylbenzamide	520	1.28	21.78	.329
			29.03	.392
$K = 2.9 \pm 0.1$	$\epsilon_C - \epsilon_I = -671$			
p-Chloro-N,N-di-			4.79	0.144
methylbenzamide	520	1.82	11.95	.306
			23.91	.493
$K = 2.8 \pm 0.1$	$\epsilon_C - \epsilon_I = -681$			
p-Chloro-N,N-di-			6.19	0.169
methylbenzamide	520	1.70	18.56	.393
			24.75	.470
$K = 2.7 \pm 0.1$	$\epsilon_C - \epsilon_I = -684$			
o-Chloro-N,N-di-			3.32	0.086
methylbenzamide	520	1.31	4.15	.104
			8.29	.189
			16.58	.311
$K = 3.1 \pm 0.1$	$\epsilon_C - \epsilon_I = -696$			
o-Chloro-N,N-di-			8.20	0.180
methylbenzamide	520	1.24	10.26	.213
			20.60	.340
			41.03	.482
$K = 3.4 \pm 0.1$	$\epsilon_C - \epsilon_I = -667$			
o-Chloro-N,N-di-			7.08	0.159
methylbenzamide	520	1.24	8.85	.188
			17.70	.307
			35.40	.450
$K = 3.4 \pm 0.1$	$\epsilon_C - \epsilon_I = -662$			
3,4-Dichloro-N,N-di-			5.85	0.116
methylbenzamide	520	1.51	14.63	.254
			29.26	.420
$K = 1.8 \pm 0.1$	$\epsilon_C - \epsilon_I = -810$			
3,4-Dichloro-N,N-di-			8.55	0.110
methylbenzamide	520	1.19	10.68	.134
			21.37	.234
			42.74	.374
$K = 1.6 \pm 0.1$	$\epsilon_C - \epsilon_I = -785$			
2,4-Dichloro-N,N-di-			6.88	0.141
methylbenzamide	520	1.51	17.21	.292
			34.42	.451
$K = 2.4 \pm 0.1$	$\epsilon_C - \epsilon_I = -669$			
2,4-Dichloro-N,N-di-			6.95	0.108
methylbenzamide	520	1.19	8.69	.131
			17.37	.227
			34.75	.355
$K = 2.1 \pm 0.1$	$\epsilon_C - \epsilon_I = -700$			
p-Methyl-N,N-di-		1.68	8.17	−0.533
methylbenzamide	450	1.68	16.34	− .825
		1.47	7.61	− .441
		1.47	22.84	− .848
$K = 5.1 \pm 0.2$	$\epsilon_C - \epsilon_I = 1077$			
p-Methyl-N,N-di-			3.15	0.134
methylbenzamide	520	1.51	3.94	.163
			7.89	.283
			15.78	.444
$K = 4.6 \pm 0.2$	$\epsilon_C - \epsilon_I = -704$			
o-Methyl-N,N-di-			3.35	0.125
methylbenzamide	520	1.31	4.19	.152
			8.39	.259
			16.78	.404
$K = 4.8 \pm 0.1$	$\epsilon_C - \epsilon_I = -693$			
p-Methoxyl-N,N-di-			4.95	0.190
methylbenzamide	520	1.32	6.19	.225
			12.37	.362
			24.75	.517
$K = 5.3 \pm 0.1$	$\epsilon_C - \epsilon_I = -686$			
p-Methoxyl-N,N-di-			3.14	0.104
methylbenzamide	520	1.08	3.93	.125
			7.86	.214
			15.71	.329
$K = 5.4 \pm 0.1$	$\epsilon_C - \epsilon_I = -665$			
o-Methoxyl-N,N-di-			4.14	0.203
methylbenzamide	520	1.31	5.18	.238
			10.36	.377
			20.72	.533
$K = 7.1 \pm 0.2$	$\epsilon_C - \epsilon_I = -679$			

[a] Error is calculated at the 90% confidence level.

TABLE II
ENTHALPIES CALCULATED AT DIFFERENT WAVE LENGTHS ASSUMING A TEMPERATURE INDEPENDENT $\epsilon_C - \epsilon_I$

Amide	Wave length, mμ	$C_I \times 10^3$	$C_D \times 10^2$	$\epsilon_C - \epsilon_I$	$-\Delta H_{calcd.}$
p-Methyl-N,N-di-	450	1.55	6.85	1137	4.5
methylbenzamide	460	1.55	6.85	1056	4.3
	476	1.55	6.85	641	3.4
	520	1.35	6.85	−704	4.9
	550	1.35	6.85	−608	4.1
N,N-Dimethyl-	450	1.74	3.78	1107	4.0
benzamide[a]	460	1.74	3.78	1045	4.4
	476	1.44	4.81	665	3.5
	520	1.71	7.39	−686	4.8
	550	1.74	3.78	−605	3.9
p-Chloro-N,N-di-	450	1.44	8.17	1083	3.7
methylbenzamide	450	1.35	15.60	1083	3.9
	450	1.35	10.40	1083	3.9
	470	1.35	15.60	805	3.3
	520	1.35	10.40	−678	4.4
	550	1.35	10.40	−606	3.9
o-Chloro-N,N-di-	450	1.55	8.74	1050	4.0
methylbenzamide	460	1.55	8.74	1002	4.2
	476	1.55	8.74	640	3.6
	520	1.35	13.13	−696	4.7
	550	1.35	13.13	−622	4.3
3,4-Dichloro-N,N-di-					
methylbenzamide	450	1.31	11.79	1203	3.5

[a] See ref. 3.

TABLE III
COMPLEX MAXIMA AND ISOSBESTIC POINTS FOR VARIOUS AMIDES AT 25°

Amide	Wave length of isosbestic point	Wave length of complex max.
Dimethylbenzamide	493	460
p-Chloro-N,N-dimethylbenzamide	493	459
o-Chloro-N,N-dimethylbenzamide	494	459
3,4-Dichloro-N,N-dimethyl-benzamide	495	459
p-Methyl-N,N-dimethylbenzamide	493	457

maximum where little percentage change in $\epsilon_C - \epsilon_I$ with temperature is found. For the dimethylbenzamide-iodine system the wave length employed was 450 mμ.

Justification for assuming that $\epsilon_C - \epsilon_I$ will be constant with temperature at 450 mμ for all of the systems studied is obtained from (a) the similarity in the wave lengths at which the isosbestic points occur, and (b) the similarity in the wave lengths of the complex maxima obtained upon crude resolution of the spectra. These values are given in Table III.

Three experiments were performed at 520 mμ in order to determine the enthalpies by a different method and to detect the change in $\epsilon_C - \epsilon_I$ with temperature at that wave length. By working with the same solution at several temperatures, the change in $\epsilon_C - \epsilon_I$ can be determined accurately from a series of such experiments. The details of this type of experiment have been dis-

TABLE IV

EQUILIBRIUM CONSTANTS OBTAINED AT 520 mμ AT DIFFERENT TEMPERATURES USING DIFFERENT AMIDE SOLUTIONS

Amide	$C_{I_2} \times 10^2$	$C_A \times 10^2$	K	$\epsilon_C - \epsilon_I$	Temp., °C.
o-Chloro-N,N-dimethylbenzamide	1.24	7.08	3.36	−662	25.0
		8.85			
		17.70			
		35.40			
	1.23	6.99	2.61	−655	35.0
		8.74			
		17.48			
		34.97			
	1.22	6.92	2.26	−623	43.3
		8.65			
		17.31			
		34.62			

$\Delta H^0 = -4.1 \pm 0.6$ kcal. mole^{-1}

	1.24	8.20	3.38	−667	25.0
		10.25			
		20.60			
		41.03			
	1.23	8.13	2.92	−659	31.0
		10.17			
		20.35			
		40.70			
	1.23	8.09	2.57	−659	35.9
		10.12			
		20.23			
		40.47			
	1.21	8.00	2.22	−624	44.8
		10.00			
		20.00			
		40.01			

$\Delta H^0 = -4.1 \pm 0.9$ kcal. mole^{-1}

p-Methoxyl-N,N-dimethylbenzamide	1.32	4.95	5.33	−686	25.0
		6.19			
		12.37			
		24.75			
	1.31	6.14	4.50	−662	31.2
		12.28			
		24.57			
	1.30	6.09	3.95	−660	37.0
		12.19			
		24.38			
	1.29	6.03	3.28	−651	44.8
		12.07			
		24.14			

$\Delta H^0 = -4.6 \pm 0.2$ kcal. mole^{-1}

cussed previously.[3] The calculated values of K and $\epsilon_C - \epsilon_I$ at different temperatures and the enthalpies obtained are given in Table IV.

The thermodynamic data obtained for the amides studied and that reported for dimethylbenzamide[3] are summarized in Table V.

In Fig. 1 Hammett's σ values for the substituent groups are plotted against log K/K_0 where K refers to the equilibrium constant for the substituted amide and K_0 is that for N,N-dimethylbenzamide.

In Fig. 2 Hammett's σ values for the substituent groups are plotted against the $-\Delta H$ values for the substituted benzamide adducts.

Discussion

Structure of the Adducts.—The formation of a 1:1 complex for all these amides is evidenced by a constant value for the equilibrium constant over an appreciable concentration range. This constancy is found even for p-methoxy-N,N-dimethylbenzamide, a molecule containing multiple potential donor sites. In this regard, the behavior of these systems is similar to that reported
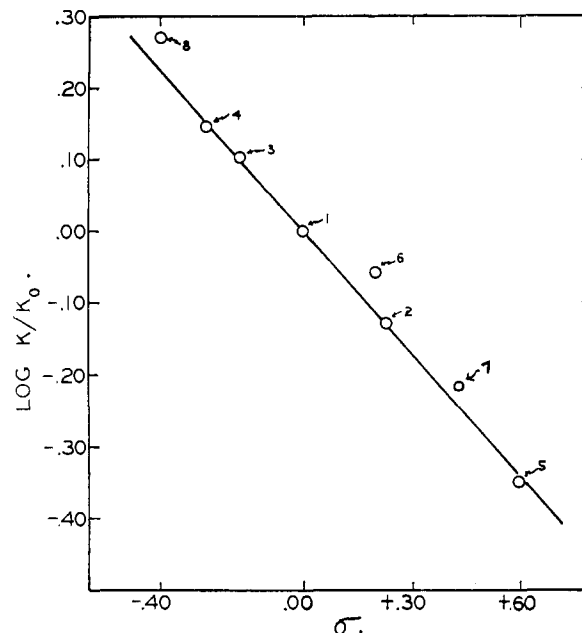

Fig. 1.—Hammett plot of log K/K_0 vs. σ values: (1) N,N-dimethylbenzamide, (2) p-chloro-N,N-dimethylbenzamide, (3) p-methyl-N,N-dimethylbenzamide, (3) o-methyl-N,N-dimethylbenzamide, (4) p-methoxyl-N,N-dimethylbenzamide, (5) 3,4-dichloro-N,N-dimethylbenzamide, (6) o-chloro-N,N-dimethylbenzamide, (7) 2,4-dichloro-N,N-dimethylbenzamide, (8) o-methoxyl-N,N-dimethylbenzamide.
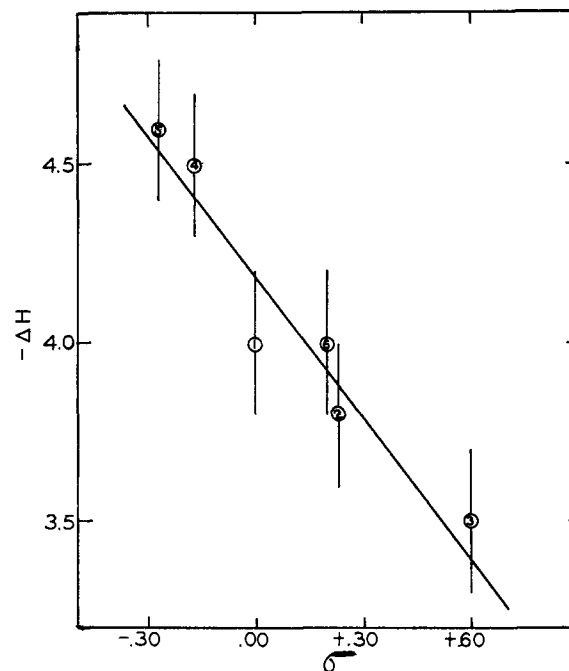

Fig. 2.—Hammett plot of $-\Delta H$ vs. σ values: (1) N,N-dimethylbenzamide, (2) p-chloro-N,N-dimethylbenzamide, (3) 3,4-dichloro-N,N-dimethylbenzamide, (4) p-methyl-N,N-dimethylbenzamide, (5) p-methoxyl-N,N-dimethylbenzamide, (6) o-chloro-N,N-dimethylbenzamide.

for dioxane, which forms only a 1:1 complex in excess dioxane.[5]

The infrared spectra in CCl₄ solution of these iodine adducts indicates that the carbonyl oxygen serves as the donor site in all the amides studied. Evidence in support of this conclusion comes from the decrease observed in the carbonyl stretching frequency of the adduct compared to the free amide. A discussion of the

(5) J. A. A. Ketelaar, *Rec. trav. chim.*, **71**, 1104 (1952).

TABLE V

SUMMARY OF THERMODYNAMIC DATA FOR THE ADDUCTS OF IODINE WITH VARIOUS SUBSTITUTED AMIDES

Amide	K_{25}	ΔF°_{25}	$-\Delta H^\circ$	$-\Delta S^\circ_{25}$
N,N-Dimethylbenzamide	3.8 ± 0.1	-0.79 ± 0.02	4.0 ± 0.2	10.7 ± 0.7
p-Chloro-N,N-dimethylbenzamide	$2.8 \pm .1$	$- .61 \pm .02$	$3.8 \pm .2$	$10.7 \pm .7$
o-Chloro-N,N-dimethylbenzamide	$3.3 \pm .1$	$- .71 \pm .02$	$4.0 \pm .2$	$11.0 \pm .7$
3,4-Dichloro-N,N-dimethylbenzamide	$1.7 \pm .1$	$- .32 \pm .03$	$3.5 \pm .2$	$10.9 \pm .6$
2,4-Dichloro-N,N-dimethylbenzamide	$2.2 \pm .1$	$- .49 \pm .03$		
p-Methyl-N,N-dimethylbenzamide	$4.8 \pm .2$	$- .93 \pm .03$	$4.5 \pm .2$	$12.0 \pm .6$
o-Methyl-N,N-dimethylbenzamide	$4.8 \pm .1$	$- .93 \pm .01$		
p-Methoxyl-N,N-dimethylbenzamide	$5.3 \pm .1$	$- .98 \pm .01$	$4.6 \pm .2$	$12.0 \pm .6$
o-Methoxyl-N,N-dimethylbenzamide	$7.1 \pm .2$	$-1.17 \pm .02$		

relationship between this shift and structure has been presented.[6]

The possible coexistence in solution of appreciable amounts of a 1:1 complex other than the oxygen coordinated species should be considered. The pi-electron density on the ring could function as a donor or in the case of the p-methoxyl amide, the ether oxygen could be a donor site. The following evidence can be presented to indicate the absence of significant (more than 5% of the complexed iodine) amounts of complexes other than that involving the carbonyl oxygen.

(1) The ring pi interaction of benzamide with iodine would be less than the benzene–iodine interaction because of the electron withdrawing effect of the amide $[-CON(CH_3)_2]$ group. At $25°$, the K for the adduct of benzene with iodine is only 0.15. Comparison of this value, an upper limit for a benzamide ring complex, with the equilibrium constants for the benzamide–iodine adducts indicates that only relatively minor amounts of a ring complex can exist.

(2) Support for the non-existence of appreciable concentrations of a ring or nitrogen complex is also obtained by the absence of an increase in the carbonyl absorption on the high frequency side of the infrared maxima.

(3) The adduct with the p-methoxyl substituted benzamide must be coördinated predominantly through the carbonyl oxygen because this K is in excellent agreement with that predicted by the σ value. If in the p-methoxyl amide there was an appreciable concentration of adduct with the ether oxygen, the net K for the two equilibria would be too large for the Hammett plot (Fig. 1). The preference of iodine for the carbonyl site instead of the ether oxygen makes a pi ring complex less feasible in all the benzamide adducts. The structure tentatively proposed for these adducts is indicated in Figure 3, where X is H or one of the substituent groups.
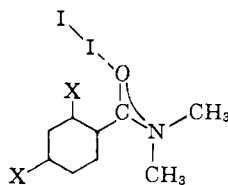

Fig. 3.—Tentative structure for the benzamide–iodine adducts.

Substituent Constant Correlations.—Both the enthalpies and equilibrium constants for the *para* substituted amides correlate very well with the Hammett σ values. Since an increased enthalpy of formation is often accompanied by an increased entropy,[7] the largest differences are encountered in the equilibrium constant values (Fig. 1). Substitution of two chlorines on the ring results in the K expected for an additive effect.

An "*ortho* effect" is apparent with the substituted benzamides as evidenced by deviations from the Ham-

mett plot shown in Fig. 1. The log K/K_0 values for the *ortho*-chloro, *ortho*-methoxyl and 2,4-dichloro substituted dimethylbenzamides lie above the line. A tentative explanation involves formation of a molecular chelate by an interaction between the chlorine or oxygen in the *ortho* position with the iodine molecule coördinated to the carbonyl oxygen (see Fig. 1). The nature of this second interaction appears to be weak and is attributed to London dispersion and electrostatic forces. The fact that the chlorine atom appears to interact nearly as strongly as the methoxyl oxygen can be explained as due to the chlorine atom being larger and more polarizable. Also the methoxyl oxygen atom is conjugated with the electron deficient benzene ring and hence its basicity is considerably reduced. It is interesting to note that the *ortho* methyl group is not large enough to give rise to a repulsive steric effect toward iodine.

A similar type molecular chelate was proposed to exist in the adducts of iodine[8] and phenol[9] with a series of alkyl substituted N,N-dimethylamides. The lack of correlation of free energy values with Taft's σ^* values was explained as being in part due to the interaction of α-chloro substituents with the iodine molecule coördinated to the carbonyl. The weakness of the interaction of the iodine molecule with the *ortho*-chloro atom is evidenced by the fact that no "*ortho* effect" can be detected by the Hammett equation for this substituent when enthalpies are plotted against σ values. The strength of this additional interaction can be placed at ≤ 0.2 kcal./mole if one assumes that the oxygen–iodine bond is not weakened by this secondary interaction.

Accuracy of the Enthalpy Procedure.—The range of the enthalpies obtained for this series is only slightly greater than one kilocalorie. This makes it imperative that the error in the enthalpy values be at a minimum. The method of determining ΔH on a single amide–iodine mixture is found to be ideally suited to a series having a narrow range of enthalpies, such as that studied here. Such a series of adducts will have similar complex absorption spectra. This suggests that the wave length where $\epsilon_C - \epsilon_I$ is constant at various temperatures will be located at the same place for the series. The absolute value of ΔH can best be obtained at the wave length where ϵ_C is independent of temperature or where any change in ϵ_C is small compared to $\epsilon_C - \epsilon_I$. The enthalpy values measured at 450 mμ satisfy these criteria and are found to obey Hammett s equation quite well. The agreement of the values at 450 mμ with those obtained at 520 mμ by a method designed to eliminate the effect of changes in molar absorptivity with temperature is added support for the accuracy of the results at 450 mμ.

(6) C. D. Schmulbach and R. S. Drago, *J. Am. Chem. Soc.*, **82**, 4484 (1960).
(7) W. B. Person, *ibid.*, **84**, 536 (1962).
(8) R. S. Drago, D. A. Wenz and R. L. Carlson, *ibid.*, **84**, 1106 (1962).
(9) M. D. Joesten and R. S. Drago, *ibid.*, **84**, 2696 (1962).